

**ICSE**

**CERTIFICATE  
CHEMISTRY**

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**X**

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**ICSE**

**CERTIFICATE**

**CHEMISTRY**

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In accordance with the latest syllabus prescribed by the Council for the Indian Certificate of Secondary Education Examination, New Delhi.

# CERTIFICATE

# CHEMISTRY

## CLASS X

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# Preface

Chemistry for Class X is designed to cover the latest Chemistry syllabus prescribed by the Council for the Indian School Certificate Examinations (ICSE), New Delhi. The Chapters are also sequenced as per the distinct directives of the Board.

We believe that Chemistry being a subject as a branch of physical science that deals with matters and its environment, it needs to be delivered in a way to attract appreciation and create a positive vibe among the readers. From the viewpoint of teachers, the contents must focus and mention every essential aspect that the students should be informed of. At the same time, the expression of the book's matter should not be markedly obscure or too much wordy that may result in the student to lose interest while reading. There has to be a proper instruction and plan for every chapter that may facilitate logically as well as comprehensive learning. Therefore, every chapter starts with 'Learning Outcomes' to highlight the expectations from that chapter that will also serve as a lesson plan to the individual teachers.

From the viewpoint of the students, the book is written with simple yet interesting language to enhance the reading experience as well as to help them understand and grasp the subject well. This will ensure that they will have a sound preparation for the examination. We deliberately avoided the book to be with only points without any explanation. This is done to support the holistic education concept of the twenty-first century and make sure that the lessons do not lose the essence of the subject.

To evoke more interest in the lessons and the subject the book is made colourful with boxes containing interesting and important information/definitions/notes. There are sections of 'in-text questions' and 'Points to Remember' to help the students quickly assess their learning outcomes from the chapters and go for a quick revision. The 'Points to Remember' section provides the synopsis of the chapters in bullet points that will help in a systematic study.

All the relevant questions as per the previous years' ICSE examinations have been incorporated in the question section after each chapter. The section also contains all the other possible questions. Moreover, this book is written to evoke interest and imagination among the students apart from helping them prepare for their examination.

We cordially invite suggestions and feedback for further improvement of this book. Positive criticism from colleagues, teachers and learned individuals, as well as students, will be highly acknowledged.

Publisher

**SYLLABUS CLASS X**  
**SCIENCE (52) : CHEMISTRY**  
**SCIENCE Paper - 2**

**Aims :**

1. To acquire the knowledge of terms, concepts, processes, techniques and principles related to the subject.
2. To develop the ability to apply the knowledge of contents and principles of chemistry in unfamiliar situations.
3. To acquire skills in proper handling of apparatus and chemicals.
4. To develop scientific temper, attitude and problem solving skills.
5. To recognize Chemical Science as having an important impact on the environment relating to cycles in nature; natural resources, pollution.

There will be one paper of **two hours** duration of 80 marks and Internal Assessment of practical work carrying 20 marks.

*The paper will be divided into two sections, Section I (40 marks) and Section II (40 marks).*

*Section I (compulsory) will contain short answer questions on the entire syllabus.*

*Section II will contain six questions. Candidates will be required to answer any four of these six questions.*

**Note :** All chemical process/reactions should be studied with reference to the reactants, products, conditions, observation, the (balanced) equation and diagram.

**1. Periodic Properties and variations of Properties – Physical and Chemical.**

- (i) Periodic properties and their variations in groups and periods.

*Definitions of following periodic properties and trends in these properties in groups and periods should be studied:*

- Atomic size,
- Metallic character
- Non-metallic character
- Ionisation potential
- Electron affinity
- Electronegativity

- (ii) Periodicity on the basis of atomic number for elements.

*Relation between atomic number for light elements (proton number) and atomic mass for light elements; the modern periodic table up to period 3 (students to be exposed to the complete modern periodic table but no questions will be asked on elements beyond period 3 – Argon); periodicity and other related properties to be described in terms of shells (not orbitals); special reference to the alkali metals and halogen groups.*

**2. Chemical Bonding**

Electrovalent, covalent and co-ordinate bonding, structures of various compounds – orbit structure and electron dot structure.

*Definition of Electrovalent Bond.*

*Structure of Electrovalent compounds NaCl, MgCl<sub>2</sub>, CaO;*

*Characteristic properties of electrovalent compounds – state of existence, melting and boiling points, conductivity (heat and electricity), ionisation in solution, dissociation in solution and in molten state to be linked with electrolysis.*

*Covalent Bond – definition and examples, structure of Covalent molecules on the basis of duplet and octet of electrons (example : hydrogen, chlorine, nitrogen, water, ammonia, carbon tetrachloride, methane.)*

*Characteristic properties of Covalent compounds – state of existence, melting and boiling points, conductivity (heat and electricity), ionisation in solution.*

*Comparison of Electrovalent and Covalent compounds.*

**Definition of Coordinate Bond :** The lone pair effect of the oxygen atom of the water molecule and the nitrogen atom of the ammonia molecule to explain the formation of  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$  ions in water and  $\text{NH}_4^+$  ion. The meaning of lone pair; the formation of hydronium ion and ammonium ion must be explained with help of electron dot diagrams.

### 3. Study of Acids, Bases and Salts

- (i) Simple definitions in terms of the molecules and their characteristic properties.  
*Self-explanatory.*
- (ii) Ions present in mineral acids, alkalis and salts and their solutions; use of litmus and pH paper to test for acidity and alkalinity.

*Examples with equation for the ionisation/dissociation of ions of acids, bases and salts: acids form hydronium ions (only positive ions) which turn blue litmus red, alkalis form hydroxyl ions (only negative ions) with water which turns red litmus blue. Salts are formed by partial or complete replacement of the hydrogen ion of an acid by a metal should be explained with suitable examples. Introduction to pH scale to test for acidity, neutrality and alkalinity by using pH paper or Universal indicator.*

- (iii) Definition of salt; types of salts.

*Types of salts : Normal salts, acid salt, basic salt, definition and examples.*

- (iv) General properties of salts :

- Deliquescence, efflorescence, water of crystallization.

Definition and example of each of the above.

- Decomposition of hydrogen carbonates, carbonates, chlorides and nitrates by appropriate acids with heating if necessary. (relevant laboratory work must be done).

*Action of dilute acids on carbonates, hydrogen carbonates and action of concentrated acid. Equations of formation of Acid rain. (Sulphuric acid) on chlorides and nitrates, to obtain carbon dioxide, hydrogen chloride and nitric acid, respectively should be taught. This will assist the students in their practical work.*

- (v) Preparation : laboratory preparation of salts (normal and acid salts) – relevant laboratory work is essential (no apparatus details are required).

*Laboratory preparation of salts (normal and acid salts): Direct combination; decomposition; displacement; double decomposition; neutralization.*

### 4. Analytical Chemistry : Use of Ammonium Hydroxide and Sodium Hydroxide

- (i) On solution of salts: colour of salt and its solution; formation and colour of hydroxide precipitated for solutions of salts of Ca, Fe, Cu, Zn and Pb; special action of ammonium hydroxide on solutions of copper salt and sodium hydroxide on ammonium salts.

*On solution of salts :*

- Colour of salt and its solution.
- Action on addition of Sodium Hydroxide to solution of Ca, Fe, Cu, Zn, and Pb salts drop by drop in excess. Formation and colour of hydroxide precipitated to be highlighted. with the help of equations.
- Action on addition of Ammonium Hydroxide to solution of Ca, Fe, Cu, Zn, and Pb salts drop by drop in excess. Formation and colour of hydroxide precipitated to be highlighted with the help of equations.
- Special action of Ammonium Hydroxide on solutions of copper salts and sodium hydroxide on ammonium salts.

- (ii) On certain metals and their oxides (relevant laboratory work is essential).

*The metals must include zinc and aluminium, their oxides and their hydroxides, which react with caustic alkalis ( $\text{NaOH}$ ,  $\text{KOH}$ ), showing the amphoteric nature of these substances.*

### 5. The Mole Concept and Stoichiometry

- (i) Gay Lussac's Law of Combining Volumes; Avogadro's Law.

*Idea of mole – a number just as dozen, a gross; Avogadro's Law - statement and explanation; Gay Lussac's Law of Combining Volumes. – statement and explanation, "the mass of 22.4 litres of any gas at S.T.P. is equal to its molar mass". (Questions will not be set on formal proof but may be taught for clear understanding) – simple calculations based on the molar volume.*

- (ii) Refer to the atomicity of hydrogen, oxygen, nitrogen and chlorine (proof not required).

*The explanation can be given using equations for the formation of HCl, NH<sub>3</sub>, and NO.*

- (iii) Relative atomic masses (atomic weight) and relative molecular masses (molecular weights): either H = 1 or 12 C = 12 will be accepted; molecular mass = 2 × vapour density (formal proof not required). Deduction of simple (empirical) and molecular formula from the percentage composition of a compound; the molar volume of a gas at S.T.P.; simple calculations based on chemical equations; both reacting weight and volumes.

*Idea of relative atomic mass and relative molecular mass – standard H atom or 1/12th of carbon 12 atom.*

*Relating mole and atomic mass; arriving at gram atomic mass and then gram atom; atomic mass is a number dealing with one atom; gram atomic mass is the mass of one mole of atoms.*

*Relating mole and molecular mass arriving at gram molecular mass and gram molecule – molecular mass is a number dealing with a molecule, gram molecular mass is the mass of one mole of molecules.*

*Molecular mass = 20 vapour density (questions will not be set on formal proof but may be taught for clear understanding); - simple calculations based on the formula.*

*Deduction of simple (empirical) and molecular formula from the percentage composition of a compound.*

## 6. Electrolysis

- (i) Electrolytes and non-electrolytes. Definitions and examples.

- (ii) Substances containing molecules only, ions only, both molecules and ions.

*Substances containing molecules only, ions only, both molecules and ions. Examples; relating their composition with their behaviour as electrolyte (strong and weak), non-electrolyte.*

*Definition and explanation of electrolysis, electrolyte, electrode, anode, cathode, anion, cation, oxidation and reduction (on the basis of loss and gain of electrons).*

- (iii) An elementary study of the migration of ions, with reference to the factors influencing selective discharge of ions, illustrated by the electrolysis of: molten lead bromide; acidified water with platinum electrodes and aqueous copper (II) sulphate with copper electrodes; electron transfer at the electrodes.

*The above electrolytic processes can be studied in terms of electrolyte used, electrodes used, ionization reaction, anode reaction, cathode reaction, use of selective discharge theory wherever applicable.*

- (iv) Applications of electrolysis: electroplating with nickel and silver; purification of copper; choice of electrolyte for electroplating.

*Reasons and conditions for electroplating;; names of the electrolytes and the electrodes used should be given. Equations for the reactions at the electrodes should be given for electroplating, refining of copper.*

- (v) Acids, bases and salts as electrolytes: reference should be made to the activity series as indicating the tendency of metals, e.g. Na, Mg, Fe, Cu, to form ions.

## 7. Metallurgy

- (i) Definition of Metals and Non-metals.

*Self-explanatory.*

- (ii) Position of the metals (alkali metals and alkaline earth metals) in the Periodic table and general characteristics applied to these elements with reference to the following – occurrence, nature, bonding, action of air, action of water, action of acids.

*Self-explanatory.*

- (iii) Comparison of Metals and Non-metals.

*General properties with special reference to physical properties: state, lustre, melting point, density, ductility, malleability, brittleness, conduction of electricity (exceptions to be specifically noted - e.g. graphite, mercury); chemical properties: a metal forms at least one basic oxide; non-metal, an acidic or neutral oxide; discharge of metallic ions at the cathode from fused metallic chlorides (link with bonding and ion formation); many metals liberate hydrogen from dilute HCl and H<sub>2</sub>SO<sub>4</sub>. In the physical properties of metals and non-metals, atomicity and valence electrons should also be included; suitable examples*

must be given for basic, acidic and neutral oxides; formation and discharge of ions at the cathode (metallic) and anode (non-metallic) should be explained with examples.

- (iv) Reduction of metallic oxides; some can be reduced by hydrogen, carbon and carbon monoxide (e.g. copper oxide, lead oxide, iron (II) oxide) and some cannot (e.g.  $\text{Al}_2\text{O}_3$ ,  $\text{MgO}$ ) - refer to activity series).

*Equations with conditions and observations should be given.*

- (v) Extraction of metals based on the activity series.

*Extraction of metals: principle of extraction of metal from its compounds by reduction – carbon reduction, electrolytic reduction. Active metals by electrolysis e.g. sodium, aluminum (reference only).*

- (vi) Corrosion of iron and its prevention.

*Experiment to illustrate that moisture and oxygen in air are responsible for the corrosion. Reaction of corrosion. Prevention by painting and galvanization.*

*Description of the changes occurring, purpose of the substances used and the main reactions with their equations.*

- (a) Uses of iron, aluminium and zinc and their alloys.

*Uses of iron, aluminium and zinc and their alloys. Composition of their alloys – steel, duralumin, brass.*

- (b) Other important alloys – bronze, fuse metal and solder.

*Uses only.*

- (vii) Metals and their alloys : common ores of iron, aluminium and zinc. Extraction of Aluminium. Metals and their alloys : Occurrence of metals in nature - mineral and ore.

*Common ores of iron, aluminium and zinc.*

*Dressing of the ore – hydrolytic method, magnetic separation, froth flotation method, chemical method by using chemical – $\text{NaOH}$  for purifying bauxite – Baeyer's Process.*

*Extraction of Aluminium: the constituents in the charge, method of electrolytic extraction*

*(flow chart to be used); structure of electrolytic cell and reason for using cryolite, electrolyte, electrodes, electrode reaction.*

*Description of the changes occurring purpose of the substances used and the main reactions with their equations.*

- (a) Uses of iron, aluminium and zinc and their alloys.

*Uses of iron, aluminium and zinc and their alloys. Composition of their alloys – steel, duralumin, brass.*

- (b) Other important alloys – bronze, fuse metal and solder.

*Uses only.*

## 8. Study of Compounds

- Hydrogen Chloride

Hydrogen chloride: preparation of hydrogen chloride from sodium chloride; refer to the density and solubility of hydrogen chloride (fountain experiment); reaction with ammonia; acidic properties of its solution.

*Preparation of hydrogen chloride from sodium chloride; (the laboratory method of preparation can be learnt in terms of reactants, product, condition, equation, diagram or setting of the apparatus, procedure, observation, precaution, collection of the gas and identification).*

*Simple experiment to show the density of the gas (Hydrogen Chloride) –heavier than air.*

*Solubility of hydrogen chloride (fountain experiment); (setting of the apparatus, procedure, observation, inference) – method of preparation of hydrochloric acid by dissolving the gas in water-the special arrangement and the mechanism by which the back suction is avoided should be learnt.*

### Reaction with ammonia

*Acidic properties of its solution - (reaction with metals, their oxides, hydroxides and carbonates to give their chlorides; decomposition of carbonates, hydrogen carbonates, sulphides, sulphites, thiosulphates and nitrates).*

### Ammonia

- (i) Ammonia : its laboratory preparation from ammonium chloride and collection; ammonia from nitrides like  $\text{Mg}_3\text{N}_2$  and  $\text{AlN}$  and ammonium salts. Manufacture by Haber's Process; density and solubility of ammonia (fountain experiment);

*aqueous solution of ammonia; its reactions with hydrogen chloride and with hot copper (II) oxide and chlorine; the burning of ammonia in oxygen; uses of ammonia.*

*Laboratory preparation from ammonium chloride and collection (the preparation can be studied in terms of, setting of the apparatus and diagram, procedure, observation, collection and identification).*

*Manufacture of ammonia on a large scale - reference should be made to Haber Process for the manufacture of ammonia.*

*Ammonia from nitrides like Mg<sub>3</sub>N<sub>2</sub> and AlN and ammonium salts; the reactions can be studied in terms of reactant, product, condition, equation.*

*Density and solubility of ammonia (fountain experiment); the property can be learnt in terms of setting of the apparatus, procedure and observation and inference.*

*Aqueous solution of ammonia - reaction with sulphuric acid, nitric acid, hydrochloric acid and solutions of iron(III) chloride, iron(II) sulphate, lead nitrate, zinc nitrate and copper sulphate.*

*Its reaction with: hydrogen chloride, hot copper (II) oxide, with chlorine in excess and ammonia in excess, burning of ammonia in oxygen; all these reactions may be studied in terms of reactants, products, condition, equation and observation; reference should be made to preparation of nitrogen from air and from ammonium nitrite.*

*Uses of ammonia - manufacture of fertilizers, explosives, nitric acid, refrigerant gas (Chlorofluoro carbon – and its suitable alternatives which are non-ozone depleting), cleansing agents, source of hydrogen.*

- (ii) The catalytic oxidation of ammonia, as the source of nitric acid; (refer to Ostwald process) simple diagram for a catalytic oxidation of ammonia in the laboratory (with conditions and reactions only).

Self-explanatory.

- Nitric Acid

*Nitric Acid : One laboratory method of preparation of nitric acid from potassium nitrate or sodium nitrate. Nitric acid as an oxidizing agent.*

*Nitric Acid : Laboratory method of preparation of nitric acid from potassium nitrate or sodium nitrate; the laboratory method can be studied in terms of reactant, product, condition, equation, setting, diagram, precaution, collection, identification.*

*As an oxidising agent: its reaction with copper, carbon, sulphur.*

- Sulphuric Acid

*Sulphuric Acid : Its behaviour as an acid when dilute, as an oxidizing agent when concentrated - oxidation of carbon and sulphur; as a dehydrating agent - dehydration of sugar and copper (II) sulphate crystals; its non-volatile nature.*

*Manufacture by Contact process (reference only). Detail of the process to be avoided.*

*Its behaviour as an acid when dilute - reaction with metal, metal oxide, metal hydroxide, metal carbonate, metal bicarbonate, metal sulphite, metal sulphide.*

*Concentrated sulphuric acid as an oxidizing agent - the oxidation of carbon and sulphur.*

*Concentrated sulphuric acid as a dehydrating agent- (a) the dehydration of sugar (b) Copper(II) sulphate crystals.*

*Non-volatile nature of sulphuric acid - reaction with sodium or potassium chloride and sodium or potassium nitrate.*

## 9. Organic Chemistry

- (i) Introduction to Organic compounds.

*Unique nature of Carbon atom – tetra valency, catenation, formation of single, double and triple bonds, straight chain, branched chain and cyclic compounds.*

- (ii) Structure and Isomerism.

*Structure of compounds with single, double and triple bonds; Isomerism – structural (chain, position)*

- (iii) Homologous series – characteristics with examples.

*Alkane, alkene, alkyne series and their gradation in properties and the relationship with the molecular mass or molecular formula.*

(iv) Simple nomenclature.

*Simple nomenclature - of the hydrocarbons with simple functional groups – (double bond, triple bond, alcoholic, ether, aldehydic, keto, carboxylic group) longest chain rule and smallest number for functional groups rule – trivial and IUPAC names.*

(v) Hydrocarbons: alkanes, alkenes, alkynes.

*Alkanes - general formula; methane (green house gas) and ethane - methods of preparation from sodium ethanoate (sodium acetate), sodium propanoate (sodium propionate), from iodomethane (methyl iodide) and bromoethane (ethyl bromide). Oxidation of methane and ethane in presence of oxygen under suitable conditions, reaction of methane and ethane with chlorine through substitution.*

*Alkenes : (unsaturated hydrocarbons with a double bond); ethene as an example. Methods of preparation of ethene by dehydro halogenation reaction and dehydration reactions.*

*Alkynes : (unsaturated hydrocarbons with a triple bond); ethyne as an example of alkyne; Methods of preparation from calcium carbide and 1,2 dibromoethane ethylene dibromide). Only main properties, particularly addition products with hydrogen and halogen namely Cl, Br and I; structural formulae of hydrocarbons. Structural formula must be given for: alkanes (up to butane), alkene ( $C_2H_4$ ); alkynes ( $C_2H_2$ ). Uses of methane, ethane, ethene, acetylene.*

(vi) Alcohols : ethanol – preparation, properties and uses.

*Preparation of ethanol :*

- *hydration of ethene;*
- *by hydrolysis of alkyl halide;*
- *Properties – Physical: Nature, Solubility, Density, Boiling Points. Chemical: Combustion, Oxidation with acidified Potassium dichromate, action with sodium, ester formation with acetic acid, dehydration with conc. Sulphuric acid with reference to Ethanol.*
- *Denatured alcohol:*
- *Important uses of Ethanol.*

(vii) Carboxylic acids (aliphatic - mono carboxylic acid): Acetic acid – preparation, properties and uses of acetic acid.

*Preparation of acetic acid from Ethyl alcohol.*

*Properties of Acetic Acid: Physical properties – odour (vinegar), glacial acetic acid (effect of sufficient cooling to produce ice like crystals). Chemical properties – action with litmus, alkalis and alcohol (idea of esterification).*

*Uses of acetic acid.*

## INTERNAL ASSESSMENT OF PRACTICAL WORK

Candidates will be asked to observe the effect of reagents and/or of heat on substances supplied to them. The exercises will be simple and may include the recognition and identification of certain gases and ions listed below. The examiners will not, however, be restricted in their choice to substances containing the listed ions.

**Gases :** Hydrogen, Oxygen, Carbon dioxide, Chlorine, Hydrogen chloride, Sulphur dioxide, Hydrogen sulphide, Ammonia, Water vapour, Nitrogen dioxide.

**Ions :** Calcium, Copper, Iron, Lead, Zinc and Ammonium, Carbonate, Chloride, Nitrate, Sulphide, Sulphite and Sulphate.

Knowledge of a formal scheme of analysis is not required. Semi-micro techniques are acceptable but candidates using such techniques may need to adapt the instructions given to suit the size of the apparatus being used.

Candidates are expected to have completed the following minimum practical work :

1. Make a solution of the unknown substance: add sodium hydroxide solution or ammonium hydroxide solution, make observations and give your deduction. Warming the mixture may be *needed*. Choose from substances containing  $Ca_2^+$ ,  $Cu_2^+$ ,  $Fe_2^+$ ,  $Fe_3^+$ ,  $Pb_2^+$ ,  $Zn_2^+$ ,  $NH_4^+$ .
2. Supply a solution of a dilute acid and alkali. Determine which is acidic and which is basic, giving two tests for each.

3. Add concentrated hydrochloric acid to each of the given substances, warm, make observations, identify any product and make deductions :  
 (a) copper oxide (b) manganese dioxide.
4. Use of pH in soil analysis, water analysis, medical field – simple identification with universal indicator.

### EVALUATION

The assignments/project work are to be evaluated by the subject teacher and by an External Examiner. (The External Examiner may be a teacher nominated by the Head of the school, who could be from the faculty, **but not teaching the subject in the section/class**. For example, a teacher of Chemistry of Class VIII may be deputed to be an External Examiner for Class X Chemistry projects.)

The Internal Examiner and the External Examiner will assess the assignments independently.

#### Award of marks (20 Marks)

Subject Teacher (Internal Examiner)	10 marks
External Examiner	10 marks

The total marks obtained out of 20 are to be sent to the Council by the Head of the school.

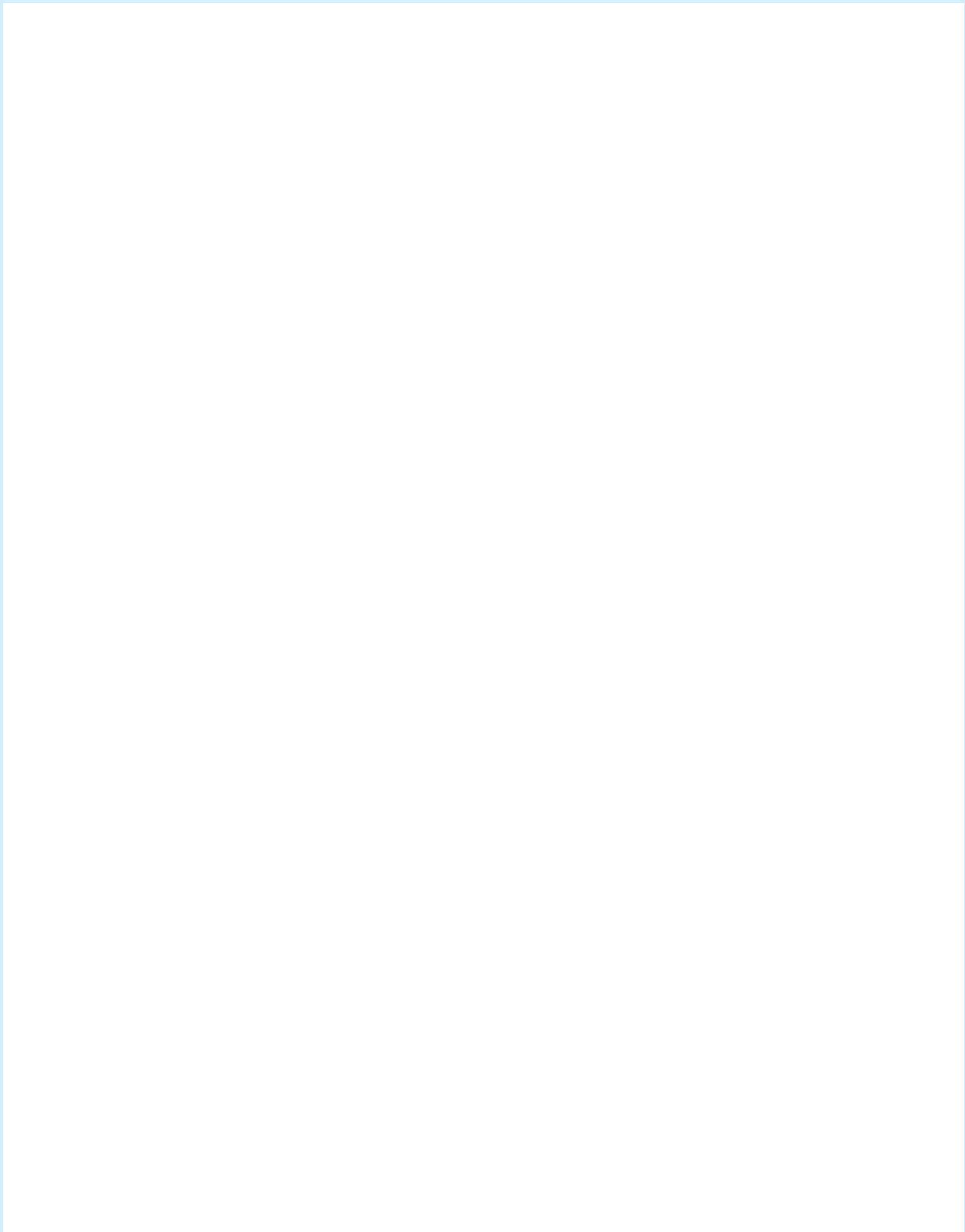
The Head of the school will be responsible for the entry of marks on the mark sheets provided by the Council.

**NOTE :** According to the recommendation of International Union of Pure and Applied Chemistry (IUPAC), the groups are numbered from 1 to 18 replacing the older notation of groups IA ..... VIIA, VIII, IB ..... VIIIB and 0. However, for the examination both notations will be accepted.

Old notation	IA	IIA	IIIB	IVB	VB	VIIB	VIII	IB	IIB	IIIA	IVA	VA	VIA	VIIA	0			
New notation	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18

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# 1



# PERIODIC PROPERTIES AND VARIATIONS OF PROPERTIES – PHYSICAL AND CHEMICAL

## LEARNING OUTCOMES

### 1.0. Introduction

#### 1.0. INTRODUCTION

All materials around us can be called as matter. Matter is everything around us in this universe that has mass (amount of material present within an object) and occupies some space. Even the atoms and molecules are all composed of matter.

On the other hand, matter can also be considered to be composed of different elements. An element is the simplest form of matter that is composed of only one type of atom that cannot be broken down further into simpler or more fundamental substances by any physical or chemical methods.

All atoms in an element are similar, whereas, for each element, the atom is unique. There are more than hundred elements known today. With increasing knowledge of elements it became a necessity to clearly classify them, in order to make things easier. To study the properties of elements, we need to arrange the elements in an orderly manner ; many attempts were made by many scientists such as Dobereiner, Newlands, Meyer and Mendeleev to arrange the elements in a periodic table.

The periodic table of elements can be considered as one of the most important and essential tools of chemistry that provides concise and fundamental information regarding every individual element

### 1.1. Modern Periodic Table

### 1.2. Periodic Properties.

present in this universe. It also gives an account about general trends across all the elements.

#### Lavoisier :

In the year 1789 Lavoisier classified the elements as metals and non-metals. Lavoisier's attempt to classify elements is not fully accepted as his classification is based on metals and non-metals only and two of his elements (heat and light) were not considered matter at all. However, his list must be acknowledged based on the 18th century perspective.

#### Dobereiner's Triad Law :

In the year 1817, Dobereiner a German scientist classified the elements having similar chemical properties.

**Fun Facts :** Elements are same everywhere. The iron (Fe) present in Earth will be same in Mars as well.

They are arranged in a group of three known as triads, such that the atomic weight of the central element is the arithmetic mean of the atomic weights of the other two elements.

For example, lithium, sodium and potassium forms a triad, the atomic weight of sodium is 23. This is the average of the atomic weights of the other two elements. Thus,

$7+39/2 = 23$ . This is satisfactory. It is following triad law.



Ca (40), S (88), and Ba (137):  $40+137/2=88.5$ . This is not following **triax law**.

Thus, this method was not satisfactory because several newly discovered elements did not fit into it.

### Periodic Table

- Shows all known elements in the universe.
- Organizes the elements by chemical properties.

### Newland's Octave Law :

The next attempt was made by an English chemist **John Newland** in 1865 based on the musical notes i.e., Every eighth element resembles the first element with respect to physical and chemical properties, like the eighth note of an octave in music. This theory was also not satisfactory. Newland's law of octaves was found to be not true in the case of elements with higher atomic weights.

### Julius Lother van Meyer :

In the year 1870, **Julius Lother van Meyer**, a German chemist plotted the atomic volumes of the elements against their atomic weight. Atomic volume is the ratio of the atomic weight to the density. A curve with several peaks was obtained with the most electropositive elements appearing at the peaks of the curves in order of their atomic weights, i.e. elements with similar properties occupy similar positions in the curve.

**Fun Facts :** The symbol of gold came from the Latin word 'aurum' that means 'shining dawn'.

### For example :

Alkali metals occupy the peak of the curve.

Halogens occur at the ascending portion of the curve. But Meyer's periodic table of elements did not find acceptance.

### Mendeleev's Periodic Law :

In the year 1871 **Mendeleev**, a Russian Chemist arranged elements based on increasing atomic weight. According to him "**the physical and chemical properties of elements are periodic functions of their atomic weights**". This is also not accepted because there was no place for isotopes in the table and no explanation for the periodicity in the properties of elements, was given.

### Henry Moseley :

The defects in Mendeleev's periodic table led scientists to determine the cause of periodicity and it was found by **Henry Moseley** in 1913, when he put forward the modern periodic table which states that **the fundamental properties of an element are periodic functions of their atomic number and not their atomic weights**.

### Bohr and Bury :

In 1921, **Bohr and Bury** gave the extended form of the table known as the long form of the periodic table or the modern periodic table.

Since the chemical properties depend on the number and arrangement of electrons in an atom, the classification of elements would be more appropriate with atomic numbers rather than atomic weights. This resulted in the development of the modern periodic table.

### 1.1. MODERN PERIODIC TABLE

#### Modern periodic law :

The law states that '**the physical and chemical properties of the elements are the periodic functions of their atomic number**'. This leads to the periodicity.

### Periodicity

The properties that appear again and again at regular intervals, or as gradual variations, i.e., increase or decrease at regular intervals, are called periodic properties.

#### How do you read the Periodic Table ?

6	←	Atomic number
C	←	Symbol
Carbon	←	Name
12.01	←	Atomic Weight

The modern periodic table is divided into 7 horizontal rows called periods and 18 vertical columns called groups.

The reason for periodicity is a recurrence of same electronic configuration.

### For example :

In group 17 all elements have seven electrons in their respective outermost shells, therefore they show similar properties, such as :

- They are coloured non-metals.
- They form negative ions.
- They are very reactive and therefore, found in combined state.
- They are only slightly soluble in water, but they dissolve much better in organic solvents like carbon disulphide, chloroform, alcohol, etc..
- Their melting and boiling points increases regularly moving down the group.
- They are good oxidising agents.
- They form hydrides with hydrogen.
- They form acidic oxides with oxygen.

It is concluded that periodicity is due to the same number of electrons in the outermost orbit of the different elements.

### Group

Vertical columns are called groups. The groups are numbered from 1 to 18. Groups contain elements with similar properties and the number of electrons present in the outermost shell determines the group.

For example sodium has atomic number 11 with electronic configuration 2, 8, 1 hence this element contains 1 electron in the valence shell it is placed in the 1<sup>st</sup> group.

Similarly magnesium has atomic number 12, electronic configuration 2, 8, 2 as its electronic configuration. Hence it is placed in the second group.

- The eighteenth group contains the inert gases from helium to radon also called the noble gases because they hardly react with other elements owing to their stable electronic configuration.
- Each group is widely separated as they are distinctly different.
- Elements of groups 1 and 2 and from groups 13-18 are called the main group or representative elements or normal elements. The elements of groups 3-12 are called transition elements. The series of elements following lanthanum and actinium are called the inner transition elements.
- Some groups are distinctly known by their family names. For example :
- Group 1 : alkali metals (Li to Fr)
- Group 2 : alkaline earth metals (Be to Ra)
- Group 16 : oxygen family (O, S, ... Po)
- Group 17 : halogen family (F, Cl, At)
- Group 18 : inert gases (He to Rn)
- Group 18 is also called the zero group.

### Period

**The number of shells present in an atom determines its period.**

- The horizontal rows of elements are called periods. There are seven periods (1-7).

Elements of period one have one shell.

Elements of period two have two shells.

Elements of period three have three shells and so on.....

If the number of the shell increases the period in the table increases. For example :

**Table 1.1**

Element with atomic numbers	No. of shells	Electronic configuration with K L M N O P	Period
F (9)	2	2,7	Second

Cl (17)	3	2,8,7	Third
Br (35)	4	2,8,18,7	Fourth
I (53)	5	2,8,18,18,7	Fifth
At (85)	6	2,8,18,32,18,7	Sixth

- The first period contains two elements, namely hydrogen and helium.
- The second period contains 8 elements from lithium to neon.
- The third period also contains 8 elements from sodium to argon.
- The fourth period contains 18 elements from potassium to krypton.
- The fifth period also contains 18 elements from rubidium to xenon.
- The sixth period is the longest period with 32 elements from caesium to radon; in order to fit this period in a table of 18 elements, 14 elements with similar properties are placed separately at the bottom of the table. This series of fourteen elements following lanthanum are called lanthanides (rare earths).
- The seventh period is a long and incomplete. This series of 14 elements with atomic numbers 89 to 103 (Th – Lr) following actinium are called actinides (radioactive elements).
- Lanthanides and actinides have similar physical and chemical properties because they belong to the same group 3, but they are placed at the bottom of the table separately because they are many in number and in these elements the valence electron enters the f shell.

There are total eighteen vertical columns called as groups.

### Bridge Elements

The elements of the second period show resemblance in properties to the elements of the next group of the third period, leading to a diagonal relationship. Such elements are called bridge elements, viz.

**Fun Facts :** Curium and meitnerium are the two elements named after women scientists Marie Curie and Lise Meitner respectively.

Li and Mg, Be and Al, B and Si.

**Table 1.2**

Group	1	2	13	14
Period 2	Li	Be	B	C
Period 3	Na	Mg	Al	Si



### Electronic Configuration

(i) **Number of Shells** : Electronic configuration of 1st group (alkali elements).

**Table 1.3**

Alkali elements	Atomic number	Electronic configuration
Li	3	2,1
Na	11	2,8,1
K	19	2,8,8,1
Rb	37	2,8,18,8,1
Cs	55	2,8,18,18,8,1

Electronic configuration of 17th group (halogen elements)

**Table 1.4**

Halogen elements	Atomic number	Electronic configuration
F	9	2,7
Cl	17	2,8,7
Br	35	2,8,18,7
I	53	2,8,18,18,7
At	85	2,8,18,32,18,7

Electronic configuration of 18th group (noble gas elements)

**Table 1.5**

Noble gas elements	Atomic number	Electronic configuration
He	2	2
Ne	10	2,8
Ar	18	2,8,8
Kr	36	2,8,18,8
Xe	54	2,8,18,18,8
Rn	86	2,8,18,32,18,8

From the above three tabular columns, we can see that 1st, 17th and 18th group elements exhibit a recurring pattern of electronic configurations.

(ii) Electrons in the outermost shell (valence)

The number of electrons present in the outermost shell are called valence electrons as the valency of elements depends on them.

### PERIOD

On moving from left to right across a period the valency (defined as the combining capacity of the element) increases up to group 4, and then it starts decreasing. Refer to Table 1.6. Firstly, the valency is increasing from Li upto C and then it starts decreasing from N to Ne and finally becomes 0 for neon. Please note that the number of valence electrons need not be the same as valency. For example, for oxygen the number of valence electrons, i.e., the number of electrons in the outer shell is 6, but its valency is 2. It means that oxygen has 2 free electrons available for bonding in a chemical reaction, while the other four electrons in the valence shell are paired and do not take part in a reaction.

**Example : Table 1.6**

Element of 2nd period	Li	Be	B	C	N	O	F	Ne
Atomic No.	3	4	5	6	7	8	9	10
Electronic Configuration	2,1	2,2	2,3	2,4	2,5	2,6	2,7	2,8
Valency	1	2	3	4	3	2	1	0

### GROUP

The number of valence electrons remain same across a group, for example in Table 1.7 we can observe that the valence electron remains 1 for Li, Na, K, Rb, Cs and Fr, elements of group 1.

**Table 1.7**

Elements	Electronic configuration	Valence electrons
Li	2,1	1
Na	2,8,1	1
K	2,8,8,1	1
Rb	2,8,18,8,1	1
Cs	2,8,18,18,8,1	1
Fr	2,8,18,32,18,8,1	1



<b>1</b>	<b>H</b>	HYDROGEN	<b>2</b>	<b>He</b>	HELIUM
<b>3</b>	<b>Li</b>	LITHIUM	<b>4</b>	<b>Be</b>	BERYLLIUM
<b>11</b>	<b>Na</b>	SODIUM	<b>12</b>	<b>Mg</b>	MAGNESIUM
<b>19</b>	<b>K</b>	POTASSIUM	<b>20</b>	<b>Ca</b>	CALCIUM
<b>37</b>	<b>Rb</b>	RUBIDIUM	<b>38</b>	<b>Sr</b>	SODIUM
<b>55</b>	<b>Cs</b>	CESIUM	<b>56</b>	<b>Ba</b>	BARIUM
<b>87</b>	<b>Fr</b>	FRANCIUM	<b>88</b>	<b>Ra</b>	RADIUM
<b>57</b>	<b>La</b>	LANTHANUM	<b>58</b>	<b>Ce</b>	CERIUM
<b>89</b>	<b>Ac</b>	ACTINIUM	<b>90</b>	<b>Th</b>	THORIUM
<b>6</b>	<b>Cr</b>	TITANIUM	<b>23</b>	<b>V</b>	VANADIUM
<b>40</b>	<b>Y</b>	YTTRIUM	<b>41</b>	<b>Nb</b>	NIOBIUM
<b>72</b>	<b>Zr</b>	ZIRCONIUM	<b>42</b>	<b>Mo</b>	MOLYBDENUM
<b>71</b>	<b>Hf</b>	HAFNIUM	<b>43</b>	<b>Tc</b>	TECHNETIUM
<b>73</b>	<b>Ta</b>	TANTALUM	<b>44</b>	<b>Fe</b>	IRON
<b>180</b>	<b>W</b>	TUNGSTEN	<b>45</b>	<b>Co</b>	COBALT
<b>183</b>	<b>Re</b>	RHENIUM	<b>46</b>	<b>Ni</b>	NICKEL
<b>190</b>	<b>Os</b>	OSMIUM	<b>47</b>	<b>Cu</b>	COPPER
<b>196</b>	<b>Pt</b>	PLATINUM	<b>48</b>	<b>Zn</b>	ZINC
<b>197</b>	<b>Hg</b>	Mercury	<b>49</b>	<b>Ga</b>	GALLIUM
<b>198</b>	<b>Au</b>	GOLD	<b>50</b>	<b>Ge</b>	GERMANIUM
<b>196</b>	<b>Pd</b>	PALLADIUM	<b>51</b>	<b>As</b>	ARSENIC
<b>192</b>	<b>Rh</b>	RHODIUM	<b>52</b>	<b>Se</b>	SELENIUM
<b>190</b>	<b>Ru</b>	RUTHENIUM	<b>53</b>	<b>Br</b>	BROMINE
<b>192</b>	<b>Rh</b>	RHODIUM	<b>54</b>	<b>Kr</b>	KRYPTON
<b>196</b>	<b>Pt</b>	RHODIUM	<b>55</b>	<b>Xe</b>	XENON
<b>198</b>	<b>Ir</b>	IRIDIUM	<b>56</b>	<b>Te</b>	TELLURIUM
<b>196</b>	<b>Os</b>	OSMIUM	<b>57</b>	<b>I</b>	IODINE
<b>192</b>	<b>Re</b>	RHENIUM	<b>58</b>	<b>Po</b>	POLONIUM
<b>192</b>	<b>W</b>	TUNGSTEN	<b>59</b>	<b>At</b>	ASTATINE
<b>190</b>	<b>Ds</b>	DARMSTADTUM	<b>60</b>	<b>Rn</b>	RADON
<b>106</b>	<b>Db</b>	DUBNIUM	<b>61</b>	<b>Fr</b>	FRANCIUM
<b>104</b>	<b>Rf</b>	RUTHERFORDIUM	<b>62</b>	<b>Ts</b>	OGANESSON
<b>103</b>	<b>Lr</b>	ACTINIDE	<b>63</b>	<b>Lv</b>	TENNESSINE
<b>89</b>	<b>Ra</b>	RADIUM	<b>64</b>	<b>Mc</b>	MUSCOVUM
<b>89</b>	<b>Fr</b>	FRANCIUM	<b>65</b>	<b>Cl</b>	FLEROVUM
<b>108</b>	<b>Db</b>	DUBNIUM	<b>66</b>	<b>Lu</b>	LUTETIUM
<b>106</b>	<b>Rf</b>	RUTHERFORDIUM	<b>67</b>	<b>Yb</b>	YTERBIUM
<b>106</b>	<b>Lr</b>	ACTINIDE	<b>68</b>	<b>Tm</b>	THULIUM
<b>89</b>	<b>Ra</b>	RADIUM	<b>69</b>	<b>Er</b>	ERBIUM
<b>89</b>	<b>Fr</b>	FRANCIUM	<b>70</b>	<b>Ho</b>	HOLMIUM
<b>106</b>	<b>Db</b>	DUBNIUM	<b>71</b>	<b>Dy</b>	DYSPROSIUM
<b>104</b>	<b>Rf</b>	RUTHERFORDIUM	<b>72</b>	<b>Tb</b>	TERBIUM
<b>103</b>	<b>Lr</b>	ACTINIDE	<b>73</b>	<b>Eu</b>	EUROPHIUM
<b>89</b>	<b>Ra</b>	RADIUM	<b>74</b>	<b>Sm</b>	SAMARIUM
<b>89</b>	<b>Fr</b>	FRANCIUM	<b>75</b>	<b>Pr</b>	NEODYMIUM
<b>104</b>	<b>Db</b>	DUBNIUM	<b>76</b>	<b>Nd</b>	PRASEODYMIUM
<b>103</b>	<b>Rf</b>	RUTHERFORDIUM	<b>77</b>	<b>Pm</b>	PROMETHIUM
<b>89</b>	<b>Fr</b>	FRANCIUM	<b>78</b>	<b>Am</b>	EUROPHIUM
<b>89</b>	<b>Fr</b>	FRANCIUM	<b>79</b>	<b>Np</b>	NEPTUNIUM
<b>105</b>	<b>Db</b>	DUBNIUM	<b>80</b>	<b>Pu</b>	PLUTONIUM
<b>106</b>	<b>Rf</b>	RUTHERFORDIUM	<b>81</b>	<b>Cm</b>	AMERICIUM
<b>106</b>	<b>Lr</b>	ACTINIDE	<b>82</b>	<b>Bk</b>	BERKELIUM
<b>89</b>	<b>Fr</b>	FRANCIUM	<b>83</b>	<b>Cf</b>	CALIFORNIUM
<b>89</b>	<b>Fr</b>	FRANCIUM	<b>84</b>	<b>Es</b>	EINSTEINIUM
<b>106</b>	<b>Db</b>	DUBNIUM	<b>85</b>	<b>Fm</b>	FERMIUM
<b>106</b>	<b>Rf</b>	RUTHERFORDIUM	<b>86</b>	<b>Md</b>	MENDELEIUM
<b>106</b>	<b>Lr</b>	ACTINIDE	<b>87</b>	<b>No</b>	NOBELIUM
<b>89</b>	<b>Fr</b>	FRANCIUM	<b>88</b>	<b>Lu</b>	LAWRENCEIUM

\* Lanthanide series

\*\* Actinide series

## 1.2 PERIODIC PROPERTIES

The periodic properties of elements depends upon the following characteristics :

- (i) Atomic size
- (ii) Ionisation energy
- (iii) Electron Affinity
- (iv) Electronegativity
- (v) Metallic and non-metallic character
- (vi) Electropositivity
- (vii) Chemical Reactivity

### (i) Atomic Size

The size of an atom is determined by its radius and the radius of an atom can be found by taking half of the distance between the nuclei of the two atoms of the same element that are connected by a bond, called the inter-nuclear distance.

For example the inter-nuclear distance between the hydrogen atoms (Fig 1.1) is 74 pm (unit-picometer). Using this method the atomic radii of nearly all the elements can be found.

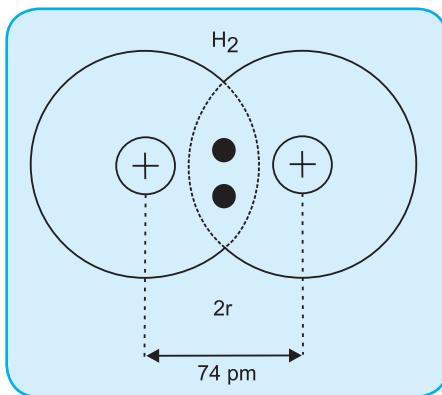


Fig. 1.1 : Inter-nuclear distance in H<sub>2</sub>

### Definition

The atomic radius of a chemical element can be defined as the measure of the size of its atoms, usually the mean or typical distance from the center of the nucleus to the boundary of the surrounding cloud of electrons or it is half the inter-nuclear distance between two like atoms which are covalently linked through a single bond, or are the nearest neighbours in a metallic solid or in rare gas in the solid state.

**UNIT :** Atomic radii are expressed in angstroms (Å) or picometers (pm)

$$\begin{aligned} \text{i.e.,} \quad 1 &= 10^{-8} \text{ cm} = 10^{-10} \text{ m} \\ 1 \text{ pm} &= 10^{-12} \text{ m} \\ 1 &= 100 \text{ pm} \end{aligned}$$

The atomic size depends on the following factors:

- number of shells
- nuclear charge

### Number of Shells

When the number of shell increases the distance between the nucleus and the outermost shell also increases and hence the atomic size increases.

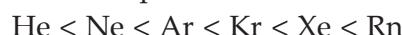
### Nuclear Charges

When the nuclear charge increases, the size of the atom decreases because when the nuclear charge increases the electrons are attracted towards the nucleus by the pull of the electrons towards the nucleus thereby decreasing the size of the atom.

Noble gases have large atomic sizes.

Noble gas atoms do not form covalent molecules and remain far apart even in the solid state. So, the inter-nuclear distance is large and the atomic size is larger than expected.

The atomic size of the inert gases increases as we move down the Group.



### Periodic Trend :

#### Group

As we move down a group the atomic size increases due to the addition of a new orbit for every element.

**For example :** If we take first group, i.e., 1 A

Table 1.8

Element with atomic no.	Atomic radius	Electronic configuration K L M N O	Size
Hydrogen (1)	37 pm	1	H
Lithium (3)	152 pm	2,1	Li
Sodium (11)	186 pm	2,8, 1	Na
Potassium (19)	231 pm	2,8, 8, 1	K ↓ Shells and atomic size increasing down a group

Li < Na < K < Rb < Cs

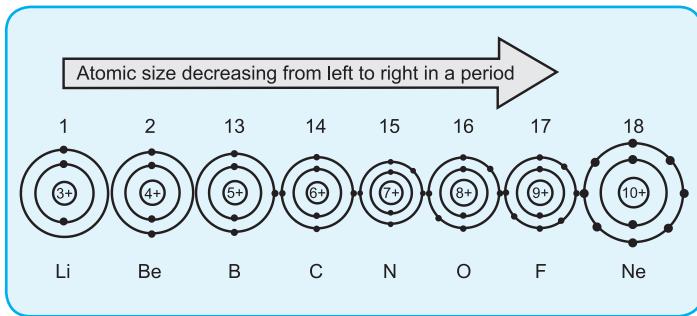
F < Cl < Br < I

In any group, the element at the top has the smallest and the element at the bottom has the biggest atom. In the groups of alkali metals and halogens, the atomic sizes vary as follows :

### Period

As we move from left to right in a period, the addition of electrons and protons increases for each atom and hence the attraction between the nucleus and the electrons increases resulting in the decrease in size of the atom.

atomic size decreasing from left to right in a period



### What are the ATOMIC NUMBER, ATOMIC WEIGHT and SYMBOL?

**Atomic Number :** The number of protons found in the nucleus of an atom or the number of electrons surrounding the nucleus of an atom.

**Atomic Weight :** The number of protons and neutrons in the nucleus of an atom.

**Symbol :** An abbreviation of the name of element.

### Ionic Size

#### Cations

A cation is formed by the removal of an electron from the outermost orbital of an atom. It is a positively charged ion, i.e., one that would be attracted to the cathode during electrolysis.

The size of the cation is small when we compare it to an atom because the nuclear charge exceeds the electronic charge. Hence, the attraction between the electrons and the nucleus increases.

For example : Na (2, 8, 1)

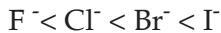


#### Anions

An anion is formed by the addition of an electron to the outermost orbit of an atom. It is a negatively charged ion, i.e., one that would be attracted to the anode during electrolysis.

The size of the anion is bigger when we compare it to an atom because as the number of electron increases without an increase in nuclear charge, the attraction between the electrons and the nucleus decreases.

For example : Cl (2, 8, 7)



#### Periodic Trend

The ionic radii increase down a group and decrease from left to right in a period. Ionic radii of alkali and halogens are as follows :

In general alkalis form cations and halogens form anions.

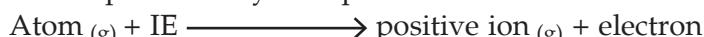
### (ii) Ionisation Energy (Ionisation Potential)

In an atom, the bonding between the nucleus and the electrons is very strong. To remove an electron from such an atom some amount of energy is required to overcome the strong attraction between the nucleus and the electrons. This is called ionisation energy (IE).

#### Definition

The minimum amount of energy required to remove the most loosely bound electron from an isolated atom in the gaseous state is known as ionisation energy.

The process may be represented as follows :



Here the gaseous atom is converted into a gaseous cation.

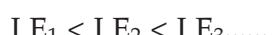
Ionisation is measured in terms of electron volts per atom (eV/atom) and its SI unit is kilo-joules per mole.

The energy required to remove one electron is known as the first ionisation energy ( $E_1$ ).

The energy required to remove the second electron from a monovalent cation is known as the second ionisation energy ( $E_2$ ).

Similarly, the energy required to remove a third electron from the divalent cation is known as the third ionisation energy ( $E_3$ ) and so on.

These ionisation energies are successive ionisation energies.



#### Factors Influencing Ionisation Energy :

##### Atomic size

As the atomic size increases, the ionisation energy decreases.

When the size of the atom increases, the distance between the nucleus and electron in the valence shell increases. So, the force of attraction between the nucleus and the electron decreases and the removal of the electron is easy. Hence the amount of ionisation energy required is less.

##### Nuclear charge

As the nuclear charge increases, the IE increases.

The greater the nuclear charge the greater is the attraction of the valence electron. Hence the valence electron is more strongly held and greater energy is required to ionise them.

#### Periodic Trend

IE decreases down a group and increases from left to right in a period.

## Group

When we move from top to bottom in a group the atomic size increases. Due to the increase in atomic size the nuclear attraction decreases. This leads to lesser IE.

**Table 1.9**

Elements	Ionisation potential
Li (2,1)	5.4
Na (2,8,1)	5.1
K (2,8,8,1)	4.3
Rb (2,8,18,8,1)	4.2
Cs (2,8,18,18,8,1)	3.9

## Period

When we move from left to right in a period the size of the atom decreases and the nuclear charge increases.

As a result, the nucleus exerts a greater force of attraction on the valence shell electrons. This leads to an increase in the ionisation potential.

IE increases from left to right in a period.

**Table 1.10**

Elements	Ionisation potential (eV)
Li (2,1)	5.4 (alkali metal)
Be (2,2)	9.3
B (2,3)	8.3
C (3,4)	11.2
N (2,5)	14.5
O (2,6)	13.6
F (2,7)	17.4 (halogen)
Ne (2,8)	21.4 (noble gas)

## Highest Ionisation Energy

IE increases up a group and from left to right in a period, the element situated at the upper right hand corner that is helium will have the highest ionisation energy.

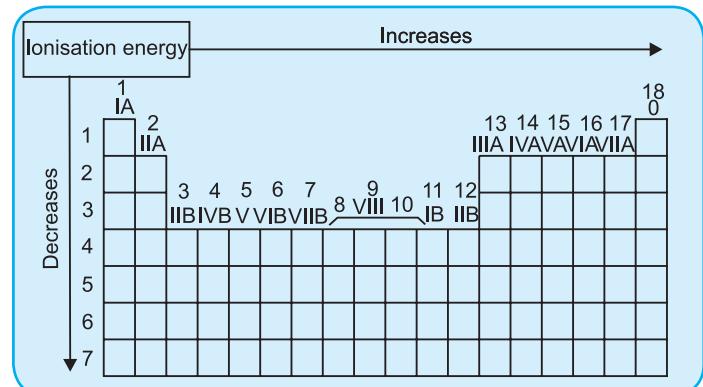
**Fun Facts :** The first element made artificially was "Technetium."

## Lowest Ionisation Energy

As the IE decreases down a group and from right to left in a period, the element situated at the lower left hand corner of the periodic table will have the lowest ionisation energy.

### NOTE

Helium will have the highest ionisation energy, while caesium (Cs) will have the lowest IE (IE of Fr is not determined correctly as it is radioactive). Metals usually have low IE whereas non-metals have high IE.

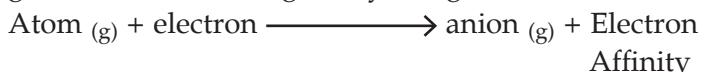


**Fig. 1.2**

## (iii) Electron Affinity

### Definition

Electron affinity (EA) or electron gain enthalpy is the amount of energy released when an isolated gaseous atom accepts an electron to form a monovalent gaseous anion or a negatively charged ion.



**Units :** Electron volt per atom (eV/atom) or KJ mole<sup>-1</sup>.

A positive electron affinity indicates that an atom is ready to lose an electron. Here the energy is absorbed.

A negative electron affinity indicates that an atom is ready to accept an electron. Here the energy is released.

### Successive Electron Affinity :

Generally, energy is released when the first electron is added to an atom, which results in the formation of a monovalent anion. Now, if another electron (second electron) is added to this anion, it experiences repulsion and energy is absorbed during this process.

Therefore, the second and further successive electron affinity values are positive.

## Factors Influencing Electron Affinity

### Atomic size

When atomic size increases, the electron affinity decreases. This is because when the atom size gets larger the attraction between the valence electron and the nucleus becomes weak. Due to this weak attraction, the electron affinity decreases and vice versa.

**Fun Facts :** There are only 90 elements found in nature. Other elements cited in a periodic table are all man-made.

### Nuclear charge

As the nuclear charge increases, so does the tendency of attracting electrons. Hence, electronegativity increases with nuclear charge.

Valence electron and the nucleus become weak due to this weak attraction and the electron affinity decreases and vice versa.

Increase in nuclear charge leads to an increase in the electron affinity because the force of nuclear attraction on the valence electrons increases. Therefore, valence electrons are more strongly held and thus a greater amount of energy is released.

#### Periodic Trend :

##### Period

Electron affinity values generally increase on moving from left to right in a periodic table. The EA values of 2nd period elements are shown below :

**Table 1.11**

Elements	Electron affinity (eV)
Li	-0.61
Be	Exception
B	-0.30
C	-1.25
N	Exception
O	-1.48
F	-3.6(Halogen)
Ne	0(noble gas)

##### Group

Electron affinity values generally decreases, when we move from top to bottom. The EA values of halogens is given below :

**Table 1.12**

Elements	Electron affinity(eV)
F	-3.60
Cl	-3.80
Br	-3.50
I	-3.20

##### Zero EA

Inert gases have zero EA. This is because the inert gas elements have completely filled sub-shells and there is no scope of adding an extra electron. Hence, they have zero electron affinity and do not form anions.

#### (iv) Electronegativity :

Electronegativity is the tendency of an atom to attract shared pair of electrons towards itself in the molecule.

Electronegativity is a dimensionless property as it is only a tendency. It only reflects the net result of the tendencies of different elements to attract the bond-forming electron pair.

Electronegativity is measured on several scales. The most widely used scale to measure electronegativity was devised by Linus Pauling.

In Pauling's scale the highest value of Electronegativity (for fluorine) is taken as 4.0 and the lowest value (for caesium) is taken as 0.7.

**Table 1.13**  
Electronegativity values in the periodic table

	H 2.1						
	Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4
Decrease in in	Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
Values	K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
	Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
	Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2

#### Factors Influencing Electronegativity

##### Atomic size

As the atomic size increases the nuclear attraction over the electrons decreases and so does the tendency of attracting electrons.

**Fun Facts :** Most of the elements in the periodic table are metals.

##### Nuclear charge

As the nuclear charge increases, so does the tendency of attracting electrons. Hence, Electronegativity increases with nuclear charge.

Hence, Electronegativity decreases with increasing atomic size.

#### Periodic Trend :

##### Period

Electronegativity increases when we move from left to right in the periodic table this is due to the increase in nuclear charge. When the nuclear charge increases the added electrons can be held more strongly.

**Table 1.14**

Elements	Electronegativity
Na	0.9
Mg	1.2
Al	1.5
Si	1.8
P	2.1
S	2.5
Ca	3.0
Ar	- (Noble gas)

## Group

Electronegativity decreases down the group in a periodic table because the atomic radius increases, pushing the valence electron farther from the nucleus.

This is because with the addition of an extra shell, the atomic size increases. Hence the nuclear charge increases and electronegativity decreases.

**Table 1.15**

Elements	Electronegativity
F	4.0
Cl	3.0
Br	2.8
I	2.5

## (vi) Electropositivity

Electropositivity is the tendency of element to donate electrons to form a positive ion (cation).

### Periodic Trend :

#### Group

Electropositivity increases when we move from top to bottom in a periodic table.

#### Period

Electropositivity decreases when we move from left to right across the periodic table.

### Difference Between Electron Affinity And Electronegativity Table 1.16

Electron affinity	Electronegativity
It is the tendency of an isolated atom to donate an electron.	It is the tendency of an atom in a molecule to attract the shared pair of electrons.
It is measured in electron volt/atom or Kcal/mol or KJ/mol.	It is a number and has no units.
It is a property of an isolated atom. An atom has an absolute value of electron affinity.	It is the property of a bonded atom. An atom has a relative value of electronegativity.
It does not change regularly in a period or a group.	It changes regularly in a period or a group.

## (v) Metallic And Non-Metallic Character

Among 118 elements, 80 are metals and remaining are non-metals.

### Metals

Metals are solids at room temperature, having a high density. They are good conductors of heat and electricity and are electropositive in nature because they have a tendency to lose their valence electrons. For example :  $\text{Na}^+$ ,  $\text{K}^+$ , etc.

## Factors Influencing Metallic Character :

### Atomic size

Larger the atomic size, the distance between the nucleus and valence electron increases and as a result the nuclear force of attraction on the valence electrons decreases. So it is easy to donate an electron from its valence shell. Hence, larger the atomic size the greater is the metallic nature.

### Nuclear charge

Greater the nuclear charge, greater is the force of attraction between the valence electrons and the nucleus. Hence, it is difficult to remove the electron from the orbit. This shows that greater the nuclear charge, lesser is the metallic nature.

### Periodic Trend :

### Group :

On moving down a group both the atomic size and nuclear charge increase. The effect of an increased atomic size is greater as compared to the increased nuclear charge. Therefore, the metallic nature increases as one moves down a group, i.e., they can lose electrons easily.

$$\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs} < \text{Fr}$$

**Example :** in group 1, lithium is the least metallic element while francium is the most metallic element.

### Period

On moving across a period, nuclear pull increases due to the increase in atomic number, and thus the atomic size decreases. Hence, the elements cannot lose electrons easily.

Therefore, the metallic nature decreases across a period, moving from left to right.

**Example :** In the 2nd period, lithium is the most metallic.

Li	Be	B	C	N	O	F
Metals		metalloids		non-metal		

In the 3rd period, sodium is the most metallic element.

### Non-Metals

Non-metals are either solids or gases at room temperature. They have low density and are bad or poor conductors of heat and electricity. They are electronegative in nature and have tendency to accept the electron to its valence shell, for example,  $\text{Cl}^-$ ,  $\text{S}^{2-}$ , etc.

Those elements which have a tendency to gain electrons, in order to attain octet in their outermost orbit, are considered as non-metals.



Non-metals usually have 5, 6, 7 electrons in their valence shell. They can gain 3, 2, 1 electron respectively, in order to attain octet structure to form an anion.



### Factors Influencing Non-Metallic Character

#### Atomic size

Smaller the atomic size, the distance between the nucleus and the valence electron is shorter and the nucleus has more pulling effect on electron. Hence the smaller the atomic size, more non-metallic is the nature of the element.

#### Nuclear charge

Greater the nuclear charge, the greater is the tendency of the atom to gain electrons. Hence the greater the nuclear charge, more non-metallic the element is in nature.

#### Periodic trend

#### Group

The atomic size increases, due to the addition of new shells over successive periods. Though the nuclear charge increases, due to an increase in atomic number, the effect of an increasing atomic size is greater, therefore, non-metallic nature decreases down the group.

Example Group 15

**Table 1.17**

Elements	Nature
Nitrogen Phosphorus	Non-metals
Arsenic	Metalloid
Antimony Bismuth	Metals

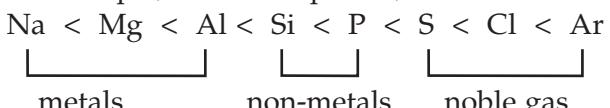
Decreasing non-metallic nature (or) increasing metallic nature

#### Period

On moving across a period, the tendency to gain electrons increases due to an increase in nuclear pull and a decrease in the atomic size.

Therefore, non-metallic character increases across a period, i.e., from left to right.

For example, in the 3rd period,



Non-metallic character increases or metallic character decreases.

#### NOTE :

The nature of oxides also shows periodicity.

In a period, the oxides of the elements shows the decrease in basic nature and finally becomes acidic.

For example, the oxides of the elements of the 3rd period.

**Table 1.18**

Elements	Nature
Na <sub>2</sub> O	Strongly basic
MgO	Basic
Al <sub>2</sub> O <sub>3</sub>	Amphoteric
SiO <sub>2</sub>	Feebly acidic
P <sub>2</sub> O <sub>5</sub>	Acidic
SO <sub>3</sub>	More acidic
Cl <sub>2</sub> O <sub>7</sub>	Most acidic

In a group the basic nature of the elements increases.

#### (vii) Chemical Reactivity

The chemical reactivity of an element depends upon their nature to lose or gain electrons from their outermost shell to attain stable electronic configuration.

The elements having 1, 2, 3 electrons in the valence shell are ready to lose electrons. Hence the chemical reactivity of metals increases.

The elements having 5, 6, 7 electrons in the valence shell are ready to gain electrons, hence the chemical reactivity of non-metals decreases.

#### Period

While moving from left to right in a period in the modern periodic table the chemical reactivity of the elements first decreases a little and then increases.

The group 1 element can lose electrons easily as compared to group 2. Similarly, group 2 elements can also lose electrons easily, but only in comparison to group 13 and not at all in comparison to group 1. As the tendency to lose electrons decreases, reactivity also decreases and thus silicon is the least reactive element in the third period. As we move from phosphorus to chlorine, the tendency to gain electrons increases, hence reactivity again increases.

**Table 1.19**

Third period elements	Reactivity
Na	Most reactive metal
Mg	Reactivity decreases
Al	Reactivity decreases
Si	Least reactive
P	Reactivity decreases
S	Reactivity decreases
Cl	Most reactive non-metal

**Group**

The tendency of loosing electrons increases down the group. Since chemical reactivity depends upon the tendency to lose electrons, it increases on going down the group.

The chemical reactivity of non-metals decreases on going down the group as it depends on the tendency to gain electrons, which decreases down the group.

**Physical Properties**

The melting and boiling points of metals decreases on moving down in a group.

**Table 1.20**

Metals	m.p.	b.p.
Li	180.5 °C	1347 °C
Na	94.5 °C	883 °C
K	63.5 °C	774 °C

The melting and boiling points of non-metals increases on moving down in a group.

**Table 1.21**

Non-metals	m.p.	b.p.	Physical state
Fluorine	-219.6 °C	-187 °C	Gas
Chlorine	-101 °C	-34.6 °C	Gas

Symbol of Element	Atomic Number Z	Mass Number A	Number of Neutrons N	Number of Protons P
H	1	1	0	1
He	2	4	2	2
Li	3	7	4	3
Be	4	9	5	4
B	5	11	6	5
C	6	12	6	6
N	7	14	7	7
O	8	16	8	8
F	9	19	10	9
Ne	10	20	10	10
Na	11	23	12	11
Mg	12	24	12	12
Al	13	27	14	13
Si	14	28	14	14
P	15	31	16	15
S	16	32	16	16
Cl	17	35,37	18,20	17,17
Ar	18	40	22	18
K	19	39	20	19
Ca	20	40	20	20

Bromine	-7.2 °C	+58.8 °C	Liquid
Iodine	+113.6 °C	+183 °C	Solid

**RELATION BETWEEN ATOMIC NUMBER AND ATOMIC MASS OF LIGHT ELEMENTS****Atomic Number (Z)**

The atomic number of an element is equal to the number of protons in the nucleus or electrons in a neutral atom.

Atomic number = number of protons = number of electrons.

**Atomic Mass or Mass Number (A)**

The Atomic mass of an element is the sum of the protons and neutrons in the nucleus of an atom.

Atomic mass = no. of proton (p) + no. of neutrons (n).

**Light Element :**

- ◆ Atomic number 1 (hydrogen) - 20 (calcium)
- ◆ They have stable nucleus

**The stable nuclei have the following characteristics :**

1. The neutron and proton numbers are even; if the proton number is odd, at least the neutron number is even.

2. If the proton number is not too large, the neutron number is the same as or close to the proton number.

**Table-1.22**

- ◆ Elements with even atomic number have equal number of protons and neutrons, which makes the

mass numbers twice the atomic number, i.e.,  $A = 2Z$ .

- ◆ Elements with odd atomic numbers except  $^1H$  and  $^{14}N$  have a mass number twice the atomic number +1 ( $A = 2Z + 1$ ) which means they have one neutron more than the number of protons.

### NOTE

Elements with  $n/p$  (neutron/proton) ratio around 1 are stable, for example, light metals like sodium, potassium, calcium etc.

Elements with  $n/p$  ratio 1.5 and above are radioactive, i.e., they emit radiations. They are unstable elements, for example, heavy metals like uranium.

	<b>Alkali metals (1st group)</b>	<b>Halogens (17th group)</b>
Elements	Li, Na, K, Rb, Cs, Fr	F, Cl, Br, I, At
Occurrence	Commonly found in a combined state, due to their reactive nature.	Commonly found in a combined state as salts.
Physical State	Shining white solid metals. Soft and can be cut with a knife. Lithium is the hardest.	Non-metals, diatomic in the gaseous state fluorine (gas) chlorine (gas), bromine (liquid), iodine (solid).
Atomic Size	They have the largest atomic size in their period (except inert gases).	They have the smallest atomic size in their period and the atomic size increases down the group.
Valency	Possess one valence electron and therefore show similar properties.	Possess seven valence electrons each and therefore show similar properties.
Conductivity	Good conductors of electricity.	Non-conductors of electricity.
Nature	Highly reactive, electropositive metals. Metallic character increases from lithium to francium.	Highly reactive, electronegative non-metals. Non-metallic character decreases from fluorine to iodine.
Melting And Boiling Point	Decreases down the group.	Increases down the group.
Ionisation Energy	They have lowest IE in their period.	They have high IE in their period.
Electron Affinity	They have low EA values which further decrease down the group.	They have high EA values that decrease down the group.

### INTEXT QUESTIONS

1. Account for the basis of Modern Periodic Table ?
2. Give Newland's law of Octave.
3. What is the basis of the Dobereimer's triad; explain ?
4. State the factors on which the Periodic properties depends.
5. Give the factors on which Ionisation energy depends.
6. Give a view of Metallic and Non-metallic character in a period and group.
7. Define atomic number, atomic weight and atomic symbol.
8. Give the relation between atomic number and atomic mass of light elements.



### POINTS TO REMEMBER

- ◆ The modern periodic law states that the properties of elements are a periodic function of their atomic number.
- ◆ Periodicity is the property which have a regular gradation along a period and down a group. There are eighteen groups and seven periods in the long form of the periodic table.



- ◆ Lanthanides and actinides are arranged separately.
- ◆ Elements of the eighteenth group are called noble gases or rare gases.
- ◆ Electropositive nature: the tendency of elements to lose electrons readily.

- ◆ Electronegative nature : the tendency of elements to attract electrons towards itself.
- ◆ Metalloids : elements with properties intermediate between those of metals and non-metals.

### EXERCISES

#### A. FILL IN THE BLANKS

1. A group of three elements with similar physical and chemical properties is called.....
2. The horizontal rows of elements in the periodic table are called.....
3. Example for an alkali metal is.....
4. Example for an inert gas is.....
5. Across a period, the ionisation potential.....
6. If the element has low ionization energy, then it is likely to be.....
7. Down the group, electron affinity.....
8. The number of electrons present in the valence shell of a halogen is.....
9. If an element has seven electrons in its outermost shell then it is likely to have the..... atomic size among all the elements in the same period.
10. In the periodic table alkali metals are placed in the group.....
11. Across a period, from left to right, electronegative character.....

#### B. MULTIPLE CHOICE TYPE QUESTIONS

Choose the Correct Answer from the options given below :

1. Dobereniers triads are groups of :
  - (a) 2
  - (b) 3
  - (c) 8
2. Iron, cobalt and nickel are placed in the.....  
  - (a) *d*-block
  - (b) I group
  - (c) Zero group
3. Zero group elements are also called.  
  - (a) Alkaline
  - (b) Noble gases
  - (c) Halogens
4. Which of the following is the periodic property ?  
  - (a) atomic weight
  - (b) atomic volume
  - (c) chaotic number
5. Find out the smallest atomic size element ?  
  - (a) Li
  - (b) H
  - (c) Na

#### C. CHOOSE THE MOST APPROPRIATE ANSWER FOR EACH OF THE FOLLOWING

- (i) Among the elements given below, the element with the least electronegativity is :  
  - (a) Lithium
  - (b) Carbon
  - (c) Boron
  - (d) Fluorine[2015]

- (ii) The metals of Group 2 from top to bottom are Be, Mg, Ca, Sr, and Ba.

- (a) Which one of these elements will form ions most readily and why?
- (b) State the common feature in the electronic configuration of all these elements. [2015]

- (iii) Arrange the following as per the instructions given in the brackets :

- (a) Cs, Na, Li, K, Rb (increasing order of metallic character).
  - (b) Mg, Cl, Na, S, Si (decreasing order of atomic size).
  - (c) Na, K, Cl, S, Si (increasing order of ionization energy)
  - (d) Cl, F, Br, I (increasing order of electron affinity)
- [2015]

#### D. MATCH THE COLUMNS

<b>Column I</b>	<b>Column II</b>
1. Lavoisier	Atomic number
2. Dobereiner	Atomic volume
3. Newland	Metals and Non-metals
4. Meyer	Triad law
5. Modern Periodic Table	Law of octave

#### E. WRITE THE BALANCED CHEMICAL EQUATIONS FOR EACH OF THE FOLLOWING

- (a) Element X is a metal with a valency 2, Y is a non-metal with a valency 3.  
  - (i) Write an equation to show how Y forms an ion.
- (b) If Y is a diatomic gas, write an equation for the direct combination of X and Y to form a compound. [2015]

#### F. VERY SHORT QUESTIONS

1. Name the first and last element in period 2.
2. Name the element which has the highest ionisation energy.
3. What is the term used for the energy released when an isolated atom in the gaseous state accepts an electron to form an anion?



4. What is the electronic configuration of the element in the third period, which gains one electron to change into an anion?
5. Give the number of the group and the period of the element having three shells with three electrons in valence shell.
6. An element has an atomic number 16.  
(a) The period to which it belongs.  
(b) The number of valence electrons.  
(c) Whether it is a metal or non-metal.
7. Which of the elements has the greatest electron affinity among the halogens?
8. Which property of an element is measured by electronegativity?

### G. SHORT QUESTIONS

1. What are the common features of the electronic configurations of elements at the end of periods 2 and 3 ?
2. How does atomic size changes from left to right in a period ?
3. What happens to the atomic size of elements on moving from top to bottom of a group ?
4. Define ionisation potential (ionisation energy).
5. State the trends in ionisation energy in  
(i) period                      (ii) group.
6. The metals of group 2 from top to bottom are; Be, Mg, Ca, Sr, Ba. Which of these metals will form ions most readily and why ?
7. How do the following properties vary in the third

- period ?  
(a) Atomic size.                      (b) Metallic character.  
(c) Ionisation energy.
8. Define the following term  
(a) Ionisation potential    (b) Electron affinity
9. An element has an atomic number 20.  
(a) The period to which it belongs.  
(b) The number of valence electrons.  
(c) Whether it is a metal or non-metal.
11. What is the common feature of the electronic configurations of the elements in group 7 ?

### H. LONG QUESTIONS

1. Discuss briefly the modern periodic table.
2. How does electronegativity of an element differ from electropositivity.
3. The elements of one short period of the periodic table are given below in order from left to right : Li, Be, B, C, O, F, Ne  
(i) To which period do these elements belong?  
(ii) One element of this period is missing. Which is the missing element and where should it be placed ?  
(iii) Place the three elements Fluorine, Beryllium and Nitrogen in the order of increasing electronegativity.  
(iv) Which one of the above elements belongs to the halogen series ?  
(v) Which one of the elements in this period shows the property of catenation ?

### RECENT BOARD QUESTIONS

#### Question 1.

- (a) Choose the correct answer from the options given below :
- (i) In the periodic table alkali metals are placed in the group..... [2011]  
(A) 1                              (B) 11  
(C) 17                              (D) 18
  - (ii) Which of the following properties do not match with elements of the halogen family ? [2011]  
(A) They have seven electrons in their valence shell.  
(B) They are highly reactive chemically.  
(C) They are metallic in nature.  
(D) They are diatomic in their molecular form.
  - (iii) An element in period-3 whose electron affinity is zero. [2012]  
(A) Neon                              (B) Sulphur  
(C) Sodium                              (D) Argon
  - (iv) An alkaline earth metal. [2012]  
(A) Potassium                              (B) Calcium  
(C) Lead                                      (D) Copper

- (v) Among the period 2 elements, the element which has high electron affinity is : [2013]  
(A) Lithium                              (B) Carbon  
(C) Chlorine                              (D) Fluorine
- (vi) Ionization Potential increases over a period from left to right because the : [2014]  
(A) Atomic radius increases and nuclear charge increases  
(B) Atomic radius decreases and nuclear charge decreases  
(C) Atomic radius increases and nuclear charge decreases  
(D) Atomic radius decreases and nuclear charge increases.
- (vii) A compound X consists of only molecules. Hence X will have : [2014]  
(A) A crystalline hard structure.  
(B) A low melting point and low boiling point.  
(C) An ionic bond.  
(D) A strong force of attraction between its molecules.



(viii) If an element A belongs to Period 3 and Group II then it will have : [2014]

- (A) 3 shells and 2 valence electrons.
- (B) 2 shells and 3 valence electrons.
- (C) 3 shells and 3 valence electrons.
- (D) 2 shells and 2 valence electrons.

(ix) Among the elements given below, the element with the least electronegativity is : [2015]

- |             |              |
|-------------|--------------|
| (A) Lithium | (B) Carbon   |
| (C) Boron   | (D) Fluorine |

(x) An element with the atomic number 19 will most likely combine chemically with the element whose atomic number is : [2016]

- |        |        |
|--------|--------|
| (A) 17 | (B) 11 |
| (C) 18 | (D) 20 |

(xi) The ratio between the number of molecules in 2g of hydrogen and 32g of oxygen is : [2016]

- |              |                            |
|--------------|----------------------------|
| (A) 1 : 2    |                            |
| (B) 1 : 0.01 |                            |
| (C) 1 : 1    |                            |
| (D) 0.01 : 1 | [Given that H = 1, O = 16] |

**(b) Give reasons for the following :**

1. Ionization potential of the element increases across a period. [2012]
2. Alkali metals are good reducing agents. [2012]
3. The oxidising power of elements increases on moving from left to right along a period in the periodic table. [2011]

**(c) Give one word or phrase for the following :**

1. The amount of energy released when an atom in the gaseous state accepts an electron to form an anion. [2014]
2. The tendency of an atom to attract electrons to itself when combined in compound. [2016]
3. The electrons present in the outermost shell of an atom. [2016]

**Question 2.**

Group number	IA 1	IIA 2	IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	0 18
2nd period	Li		D			O	J	Ne
	A	Mg	E	Si		H	M	
	R	T	I		Q	u		y

- In this table H does not represent hydrogen.
- Some elements are given in their own symbol and position in the periodic table.
- While others are shown with a letter.

**With reference to the table answer the following questions :** [2013]

1. Identify the most electronegative element.
2. Identify the most reactive element of group 1.
3. Identify the element from period 3 with least atomic size.
4. How many valence electrons are present in Q ?
5. Which element from group 2 would have the least ionization energy ?
6. (a) Identify the noble gas of the fourth period.  
(b) In the compound between A and H what type of bond would be formed and give the molecular formula for the same.
7. **Using the information above, complete the following :** [2013]

- (i) \_\_\_\_\_ is the metallic element.
- (ii) Metal atoms tend to have a maximum of \_\_\_\_\_ electrons in the outermost energy level.
- (iii) Non-metallic elements tend to form \_\_\_\_\_ oxides while metals tend to form \_\_\_\_\_ oxides.
- (iv) Non-metallic elements tend to be \_\_\_\_\_ conductors of heat and electricity.
- (v) Metals tend to \_\_\_\_\_ electrons and act as \_\_\_\_\_ agents in their reactions with elements and compounds.

	X	Y
Normal Electronic Configuration	2,8,7	2,8,2
Nature of oxide	Dissolves in water and turns blue litmus red	Very low solubility in water. Dissolves in hydrochloric acid
Tendency for oxidising and reducing reactions	Tends to oxidise elements and compounds	Tends to act as a reducing agent
Electrical and Thermal conductivity	Very poor electrical conductor Poor thermal conductivity	Good Electrical conductor Good Thermal conductor
Tendency to form alloys and amalgams	No tendency to form alloys	Forms alloys

**Question 3.**

(a) Use the letters only written in the Periodic table given below to answer the questions that follow :  
[2016]

	I	II	Groups							0
Periods			III	IV	V	VI	VII	O	L	
1			E	G	J	Z	M			
2	Q									
3	R									
4	T									
5										

1. State the number of valence electrons in atom J.
2. Which element shown forms ions with a single negative charge ?
3. Which metallic element is more reactive than R ?
3. Which element has its electrons arranged in four shells ?



# 2

# CHEMICAL BONDING

## LEARNING OUTCOMES

- 2.0. Introduction
- 2.1. Types of Chemical Bonding
- 2.2. Electrovalent or Ionic Bond

## 2.3. Lewis Dot Structure

- 2.4. Covalent Bond
- 2.5. Difference between Ionic and Covalent Bond
- 2.6. Redox Reaction
- 2.7. Coordinate Bond

## 2.0 INTRODUCTION

Water is made up of water molecules and each water molecule is made up of two hydrogen atoms and one oxygen atom that are bonded together in a stable arrangement.

For a molecule, attaining a stable arrangement means achieving the electronic configuration of the nearest noble gas, that is, its outer shell should get filled with electrons. Molecules are more stable than the atoms present within them in the native form and stable molecules are less prone to react in a chemical reaction. By nature, chemical reactions often takes place so that the atoms can achieve a more stable configuration.

*Molecules are made up of sub-units called atoms, that means that every molecule is made up of two or more atoms bonded, or linked, together.*

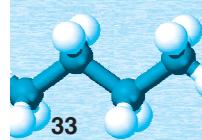
In the year 1916, W. Kossel and G. N. Lewis developed the theory of chemical bonding to explain why atoms combined to form molecules. According to this theory of chemical bonding, a chemical bond is said to be formed when atoms interact by losing, gaining or sharing valence electrons to attain a stable octet electronic configuration.

*"The union of two or more atoms through the redistribution of electrons in their outermost shell either by transference of electrons or by the sharing of electrons is called chemical bonding."*

In the periodic table only the noble gases, i.e., Ne, Ar, Kr, Xe and Rn do not react to form bonds. This is because they are stable elements by nature, i.e., their valence shell is already completely filled with eight electrons. This is why we call it as **octet configuration** with the exception of element helium which has only two electrons in the valence shell. We call this as a **duplet configuration**. Therefore, the atoms having 8 electrons (or 2 electrons) in their valence shells are stable and thus remain unreactive.

**Table 2.1 Electronic Configuration of Inert Gas**

Inert Gas	Atomic No.	K	L	M	N	O	P
He	2	2					
Ne	10	2	8				
Ar	18	2	8	8			
Kr	36	2	8	18	8		
Xe	54	2	8	18	18	8	
Rn	86	2	8	18	32	18	8

**Table 2.2. Differences between atoms and ions**

Atoms	Ions
Electrically neutral	Electrically charged, either positively or negatively
Contains equal number of protons and electrons	Contains less electrons than protons
Cannot exist independently	Can exist independently in a solution
Reactive properties, reacts with water	Reactive properties, stable in water

**The valence shell is the outermost shell of an atom and the electrons present in this shell are known as valence electrons.**

## 2.1. TYPES OF CHEMICAL BONDING

There are six types of chemical bonds :

- 1. Electrovalent or ionic bond
  - 2. Covalent bond
  - 3. Coordinate bond
  - 4. Metallic bond
  - 5. Hydrogen bond
  - 6. Vander walls bond
- }      Strong Bonds      }      Weak Bonds

The first four types of bonds are called as **strong bonds**, whereas the last two types of bonds are known as **weak bonds**. Let's discuss electrovalent or ionic bond, covalent bond and coordinate bond briefly in this chapter.

## 2.2. ELECTROVALENT OR IONIC BOND

The bond formed due to electrostatic attraction or force existing between the oppositely charged ions (cations and anions) produced by the electron transfer from one atom to the other is known as the **ionic or electrovalent bond**. Here the more electronegative element gains electron and the less electronegative element lose electrons resulting in the formation of anions and cations respectively. Usually an ionic bond is formed between a metal and a non-metal. The metal loses electrons to become a **cation** and a non-metal gains electrons to become an **anion**.

A metallic element, whose atoms readily lose electrons to form positively charged ions, are electropositive elements, e.g., sodium.



A non-metallic element, whose atoms readily accepts electrons to form negatively charged ions, are electronegative elements, e.g., chlorine.



Here the **Na<sup>+</sup> (cation)** and the **Cl<sup>-</sup> (anion)** are oppositely charged and hence attract each other to form an electrovalent bond.

Thus, the electrovalent bond is defined as :

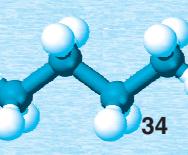
*"A chemical bond formed by the transfer of one or more electrons from one atom to another."*

The **electrovalency** of an element is the number of electrons an atom, an element loses or gains while forming an ionic bond. So, in an ionic compound the electrical charges are balanced.

For example, the metal sodium (Na) combines with non-metal chlorine (Cl) to form an ionic compound NaCl or common salt. The electrovalency of Na is +1 and that of chlorine is -1.

### Characteristics of an ionic bond

1. Ionic bonds are formed on the basis of electrostatic force of attraction between cations and anions.
2. The ionic bond is non-directional and extends in all directions. Therefore in solid state, a single ionic molecule does not exist as such. This implies that there is no intrinsically preferred direction in which a neighbour should lie for the strength of bonding to be maximized.
3. Ionic compounds do not dissolve in non-polar, organic solvents like ether, benzene, acetone, alcohol, CCl<sub>4</sub>, etc., as they have no net charge. Ionic compounds dissolve in polar solvents like water, as polar solutions consist of molecules that have a difference in charge across the molecule. Water has a high dielectric constant which weakens the electrostatic force of attraction existing between the constituent ions of the ionic solid.
4. Ionic compounds have high melting and boiling points. The strong electrostatic force of attraction between the oppositely charged ions makes the ionic bonds strong, requiring high amount of heat energy to break them.
5. Since ionic compounds are closely packed, they exhibit decreased volume and increased density.
6. Ionic compounds give ions in aqueous solutions; these ions rapidly unite and form compounds. Hence the reaction rate is greater in ionic compounds in aqueous solutions.



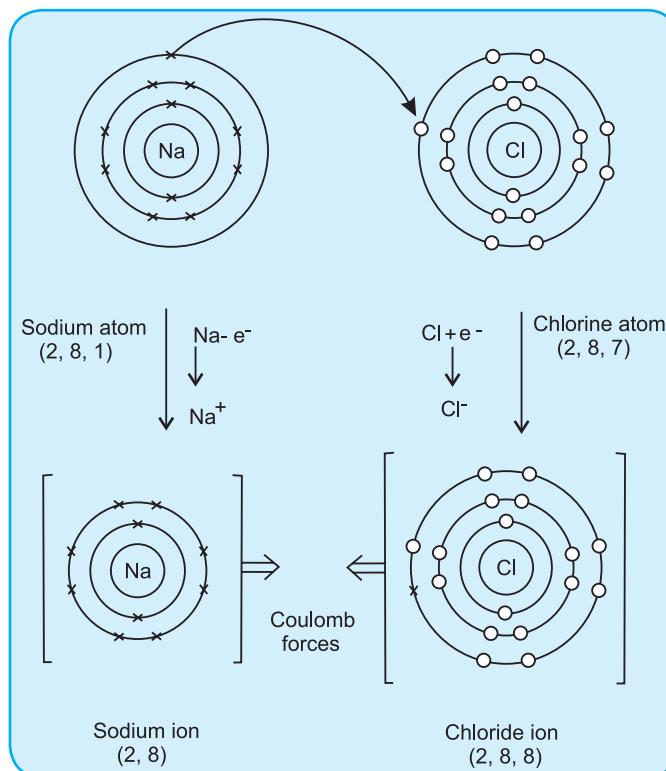
7. Due to strong coulombic forces of attraction between the oppositely charged ions, electrovalent compounds exist mostly as crystalline solids that are hard and brittle. The hardness and high lattice enthalpy contribute to the low volatility; hence they have high melting and boiling points.
8. Because of the strong electrostatic forces, the ions in the solid are not free to move and hence electrovalent compounds act as a poor conductor of electricity in the solid state. However, in the molten state, or in a solution, due to the mobility of the ions they become good conductors of electricity.
9. Some examples of electrovalent (ionic) compounds

NaCl	(Sodium chloride)
MgCl <sub>2</sub>	(Magnesium chloride)
CaO	(Calcium oxide)
KBr	(Potassium bromide)
CaCl <sub>2</sub>	(Calcium chloride)

### Ionic bond formation

#### Formation of Sodium Chloride

The ionic bond is formed by the strong electrostatic force of attraction between the ions.



**Fig. 2.1 : Structure of Sodium Chloride by Electrovalent Linkage**

The electronic configuration of sodium atom is  $2, 8, 1$  and the electronic configuration of chlorine atom is  $2, 8, 7$ ; from the above figure the sodium atom has

1 electron in its valence shell while the chlorine atom has 7 electrons in its valence shell. Therefore, to attain stable octet configuration, sodium atom transfers its valence electron (electron in the outermost shell) to the chlorine atom. This results in the formation of a cation. By giving up one electron, the sodium atom has 10 electrons and 11 protons thus has a net +1 positive charge. This positive charged atom is called a sodium ion and written as  $\text{Na}^+$ . On the other hand, by accepting one electron, the chlorine atom has 18 electrons and 17 protons; therefore, it has a net -1 negative charge. This negatively charged atom is called chlorine ion and written as  $\text{Cl}^-$  anion. The  $\text{Na}^+$  and  $\text{Cl}^-$  ions together form a sodium chloride molecule.

The two oppositely charged ions are attracted towards each other due to the electrostatic force or Coulombic force of attraction to form the ionic bond.

### 2.3 LEWIS DOT STRUCTURE

G. N. Lewis in 1919 suggested that atoms can attain inert gas configuration by sharing one or more electron pairs with similar or dissimilar atoms. By sharing the valence electron both these atoms are brought together, resulting in the formation of a covalent bond.

Lewis structures (also known as Lewis dot diagrams, electron dot diagrams, Lewis dot formulas, Lewis dot structures, and electron dot structures) are diagrams that show the bonding between atoms of a molecule and the lone pairs of electrons that may exist in the molecule.

The Lewis dot structure tells us the location of all valence electrons in a molecule including both the shared electrons in bonds as well as the non-bonded electron pairs.

Lewis dot structure has the valence electrons surrounded on the atom and they are represented by the dots on the atom. We can place four dots one by one around the atom after that we can place second dot again on the four sides.

If an atom has 1 electron on one side we call that electron as 1 unpaired electron, whereas, if an atom has 2 electrons on two sides of the atom we call the electrons as two unpaired electrons and so on.

If an atom has 2 electrons on one side and three electrons on the remaining side, we call the electrons as one paired electron and three unpaired electrons.

The electron dot symbol for an atom consists of the element symbol surrounded by the outermost shell electron and the electrons are represented by the dots symbol.

The paired electrons are represented by a pair of dots, and the unpaired electron in the outermost orbit is represented by a single dot.

**Table 2.3. Electron Dot Symbols**

Element	Symbol
Hydrogen	H•
Oxygen	:O:
Chlorine	:Cl:

Electrons can be represented as circles or crosses for the different atoms in a molecule.

For example the sodium atom is represented by the symbol Na surrounded by a white dot that represents its valence electron and the chlorine atom is represented by symbol Cl surrounded by seven black dots in the below figure.

#### The electron dot structure of NaCl



The electron from a sodium atom moves to the chlorine atom to form sodium ion (cation) and chlorine ion (anion).

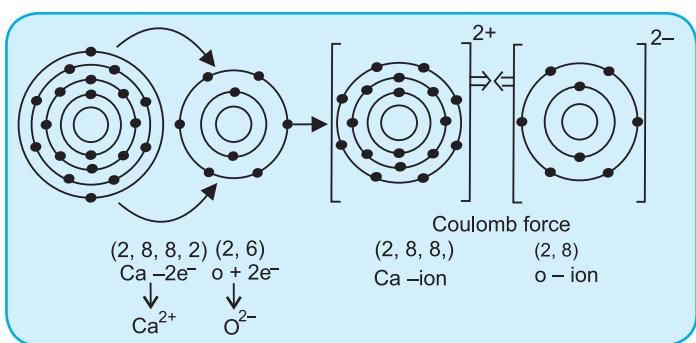
**Table 2.4. Properties of Sodium Atom and Sodium Ion**

Property	Sodium atom Na	Sodium ion Na <sup>+</sup>
Colour	Silvery white	Colourless
Toxicity	Poisonous	Non-poisonous
Chemical action	Very active	Inactive
Valence shell	Incomplete outermost shell	Complete outermost shell
Electrical state	Neutral	Positively charged
Existence	Combined state	Independent existence

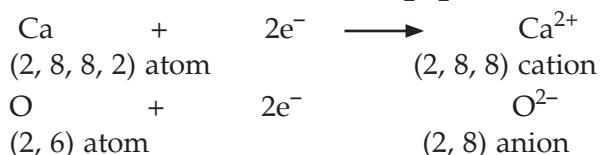
**Table 2.5. Properties of chlorine atom and chlorine ion**

Property	Chlorine atom (Cl)	Chloride ion Cl <sup>-</sup>
Colour	Yellowish green (gas)	Colourless
Toxicity	Poisonous	Non-poisonous
Odour	Suffocating	Odourless
Chemical action	Very active	Inactive
Valence shell	Incomplete outermost shell	Complete outermost shell
Electrical state	Neutral	Negatively charged
Existence	Not independent	Independent

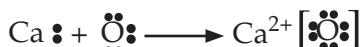
#### Formation of calcium oxide (CaO)

**Fig. 2.2 : Formation of Calcium Oxide**

The electronic configuration of Ca atom is 2,8,8,2 and the electronic configuration of oxygen atom is 2,6. Figure 2.2 shows that the calcium atom has 2 electrons in its valence shell and the oxygen atom has 6 electrons in its valence shell. Therefore, to attain stable octet configuration, the calcium atom transfers its 2 electrons to the oxygen atom. This results in the formation of Ca<sup>2+</sup> cation and O<sup>2-</sup> anion respectively, and a calcium oxide molecule is formed. Since, only one oxygen atom is needed to accept the 2 valence electrons donated by a calcium atom, the formula of calcium oxide is CaO and not Ca<sub>2</sub>O<sub>2</sub>.

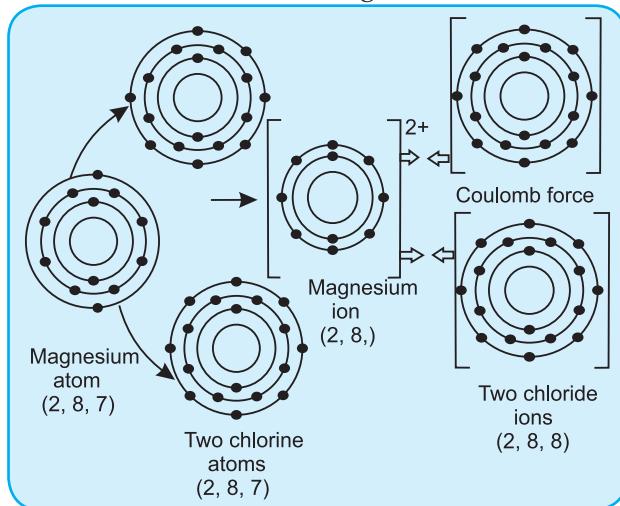


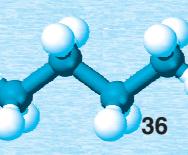
#### The Electron Dot Structure of Calcium Oxide



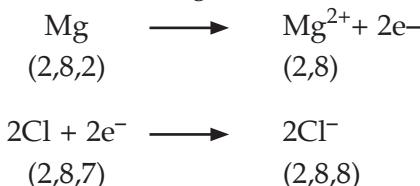
#### Formation of Magnesium Chloride (MgCl<sub>2</sub>)

The electronic configuration of magnesium is 2,8,2 and the electronic configuration of two chlorine atoms is 2,8 and 7 from the figure 2.3.

**Fig. 2.3 : Formation of Magnesium Chloride**

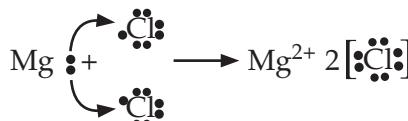


We can understand that the magnesium atom has 2 electrons in its valence shell and the two chlorine atoms have 7 electrons each in their valence shells. So, to attain octet the magnesium atom should give 1 electron to one chlorine atom and another electron to another chlorine atom and the two chlorine atoms should gain one electron each in the valence shell. This results in the formation of  $Mg^{2+}$  cation and  $2Cl^-$  anions respectively and forms the magnesium chloride molecule ( $MgCl_2$ ).



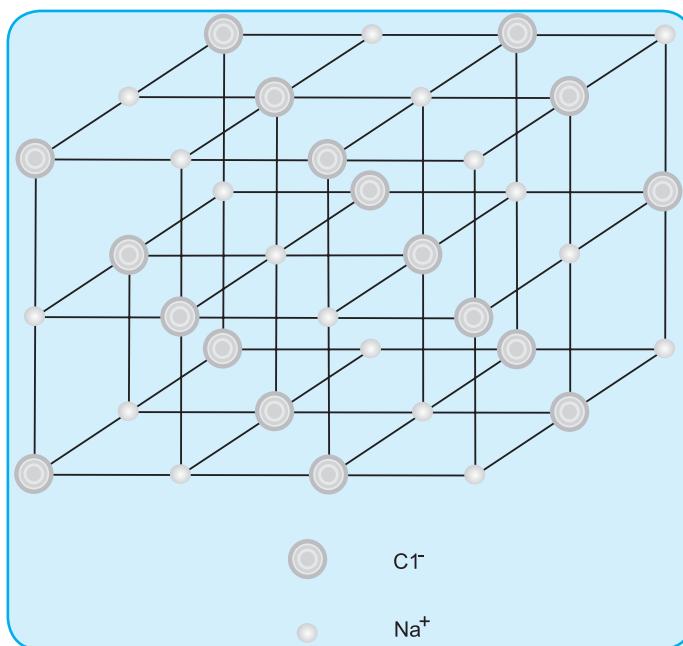
**Fig. 2.3 : Formation of Magnesium Chloride**

#### The Electron dot Structure of Magnesium Chloride



#### Structure of sodium Chloride

The sodium chloride crystal is made up of a sodium ion and a chloride ion in a three dimensional geometric pattern called a lattice.



**Fig. 2.4 : Structure of Sodium Chloride**

(See Figure 2.4). In this lattice each  $Na^+$  ion is surrounded by six  $Cl^-$  ions and each  $Cl^-$  ion is surrounded by six  $Na^+$  ions. The number of ions surrounding the oppositely charged ions is

determined by the relative size and the charge of the ions.

The ionic compound cannot conduct electricity in solid state as the ions present in the sodium chloride (solid state) are unable to move. But, in solution or in aqueous form it will conduct electricity because of the mobility of ions.

#### 2.4. COVALENT BOND

A covalent bond is a chemical bond formed when two atoms mutually share a pair of electrons. By doing so, the atoms attain stable octet electronic configuration. In covalent bonding, overlapping of the atomic orbitals having an electron from each of the two atoms of the bond takes place resulting in equal sharing of the pair of electrons. The inner-atomic bond thus formed due to the overlapping of atomic orbitals of the electrons is also known as a covalent bond. Generally, the orbits of the electrons in the valence shell of the atoms are used for electron sharing.

The shared pair of electrons lies in the middle of the covalent bond. The atoms of the covalent bond attain the stable octet configuration including the shared pair of electrons.

The covalent bond can be formed by two atoms of the same element (homonuclear diatomic molecule) or two atoms of different elements (heteronuclear diatomic molecule).

When a non-metal combines with another non-metal, sharing of electron takes place between their atoms and a covalent bond is formed, e.g.,  $HCl$ . Similarly, non-metal with the same non-metal atom combination also takes place to form a covalent bond, e.g.,  $Cl_2$  molecule.

#### Definition

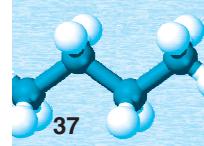
*A chemical bond formed due to the mutual sharing of electron pairs, between atoms taking an equal number of electrons from each atom is called a covalent bond and the compound formed due to this bond is called a covalent molecule.*

#### Types of covalent bond :

1. Single covalent bond
2. Double covalent bond
3. Triple covalent bond

#### 1. Single Covalent Bond

A single covalent bond is denoted by a single short line (—) between the two atoms and they



are formed by the sharing of one pair of electrons between the atoms.

e.g., hydrogen, chlorine, water, hydrogen chloride, methane, ammonia and carbon tetrachloride are examples of single covalent bond molecules.

**Table 2.6. Structures of molecules**

Molecule	Structure
Hydrogen	H — H
Chlorine	Cl — Cl
Hydrogen chloride	H — Cl
Ammonia	H — N — H H
Methane	H H — C — H H

## 2. Double Covalent Bond

A double covalent bond is denoted by a double short line (==) between the two atoms and they are formed by the sharing of two pairs of electrons between the atoms.

e.g., The oxygen molecule contains a double bond between two atoms and it can be represented as O==O

## 3. Triple Covalent Bond

A triple covalent bond is denoted by a triple short line (≡) between the two atoms and they are formed by the sharing of three pairs of electrons between the atoms.

e.g., nitrogen, ethene molecule



Some molecules have a combination of single bond as well as a double bonds, e.g., ethene ( $\text{C}_2\text{H}_4$ ).

Here we have four single bonds and one double bond.

Some molecules have a combination of single bonds and triple bond, e.g., ethyne ( $\text{C}_2\text{H}_2$ ).

The covalency of an atom is the number of its electrons taking part in the formation of shared pairs. Thus, the covalency of hydrogen is 1, oxygen is 2, nitrogen is 3 and carbon is 4.

### Covalent bonds can be polar or nonpolar

A covalent bond formed between two similar atoms or dissimilar atoms which have little or zero electronegativity difference is called a non-polar

covalent bond and the compound thus formed is a non-polar compound. These compounds do not ionize in water due to lack of charge separation.

In diatomic molecules like hydrogen, chlorine and oxygen the bonding electron pair constituting the covalent bond is equally shared by both the linked atoms. Hence these compounds have no polarity and are called as non-polar compounds.

The bond formed between dissimilar atoms can be non-polar if their electronegativity difference is less and their structure permits the shared pair of electrons to attract equally. For example, the linked atoms, like methane, carbon tetrachloride, etc.

The shared electron pairs in methane are at equal distance from the carbon and the hydrogen atoms, because the two have nearly equal electronegativities: carbon = 2.5, hydrogen = 2.1.

The covalent compounds are said to be polar when the shared pair of electrons are not at equal distance between the two atoms. This results in the development of fractional positive and negative charges on the atoms. These compounds ionize in water. For example : hydrogen chloride.

In hydrogen chloride, the strong nuclear charge of the chlorine atom (the electronegativity of chlorine is 3) attracts the shared electron pair towards itself, i.e., negative charge shifts towards the chlorine atom, thereby developing a fractional negative charge on it, whereas, the hydrogen atom develops a fractional positive charge. Therefore, a polar covalent bond is formed.

**Note :** *The more the electronegativity difference between the two atoms forming a bond, the more is the polar nature of the molecule.*

The type of bond formed between two atoms :

1. Same or similar electronegativity is non-polar
2. Slightly different electronegativity is polar
3. Much electronegativity difference is ionic

Since a polar covalent molecule has both positive and negative poles, it is also known as 'dipole molecule'.

### Formation of the hydrogen molecule ( $\text{H}_2$ )

Hydrogen molecule ( $\text{H}_2$ ) is formed between two atoms of hydrogen. These two hydrogen atoms contribute one electron each and they mutually share the contributed electrons. This can be represented by the following Lewis dot structure.

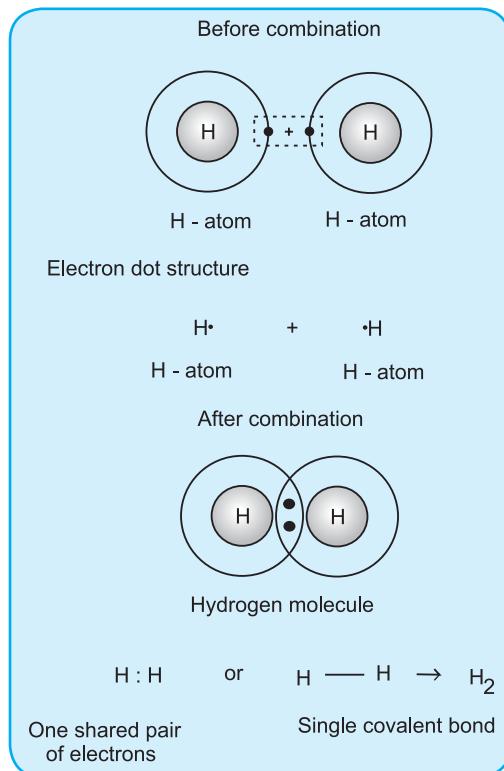
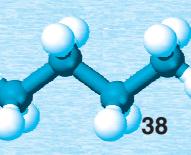


Fig. 2.5 : Formation of Hydrogen Molecule

Hydrogen has one electron in its valence shell. (See Figure 2.5.) Here we can find one pair of electrons is shared. So, it is called a single covalent bond. It is represented by one (—) symbol between the atoms sharing the electron. Both Hydrogen atoms attain the stable configuration of Helium gas, i.e., they attain duplet configuration.

#### Formation of the oxygen molecule ( $\text{O}_2$ )

Oxygen has the electronic configuration of 2, 6. (See Figure 2.6.) An oxygen molecule is formed by the combination of two oxygen atoms.

#### Electron Dot Structure



Two shared pair of electrons

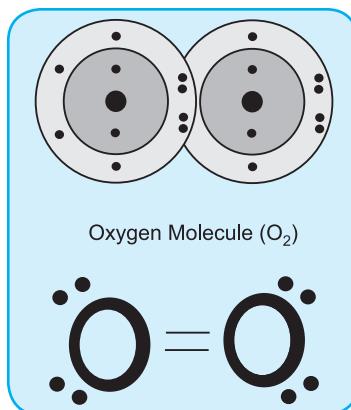


Fig. 2.6 : Formation of Oxygen Molecule

Each oxygen atom needs 2 electrons to attain octet configuration. Hence, two pairs of electrons are shared with each other. Due to this sharing a double covalent bond is formed and it has two lone pairs of electron on each atom. It is represented by (= $=$ ) symbol between the atoms.

#### Formation of the nitrogen molecule ( $\text{N}_2$ )

Nitrogen has a electronic configuration of 2, 5. (See Figure 2.7). A nitrogen molecule is formed by the combination of two nitrogen atoms.

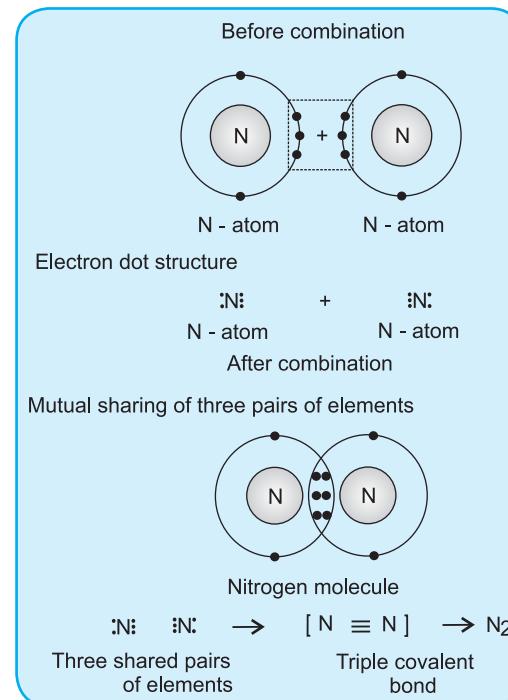


Fig. 2.7 : Formation of Nitrogen Molecule

Each nitrogen atom needs three electrons to attain stable octet configuration, hence three pairs of electrons are shared between the two nitrogen atoms. Due to this sharing a triple covalent bond is formed and it is represented by the (= $=$ ) symbol between the atoms and it has one lone pair of electrons on each atom.

#### Formation of the hydrogen chloride molecule (HCl)

The hydrogen chloride molecule is formed by the sharing of an electron pair between unlike atoms (i.e., H atom and Cl atom)(see Figure 2.8).

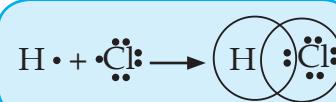
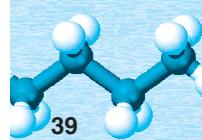


Fig. 2.8 : Formation of Hydrogen Chloride Molecule



A single covalent bond is formed. Here, hydrogen atom has one electron in the valence shell. It needs another 1 electron to attain the stable duplet configuration and the chlorine atom has 7 electrons in the valence shell and it needs 1 electron to attain stable octet configuration. Thus, both the atoms share one electron between themselves and one pair of sharing electron is formed. Due to this sharing, single covalent bond is formed and it is represented by (—) symbol between the atom.

#### Formation of the ammonia molecule ( $\text{NH}_3$ )

Nitrogen has 5 electrons in its valence shell. It needs 3 electrons to attain the octet configuration (see figure 2.9).

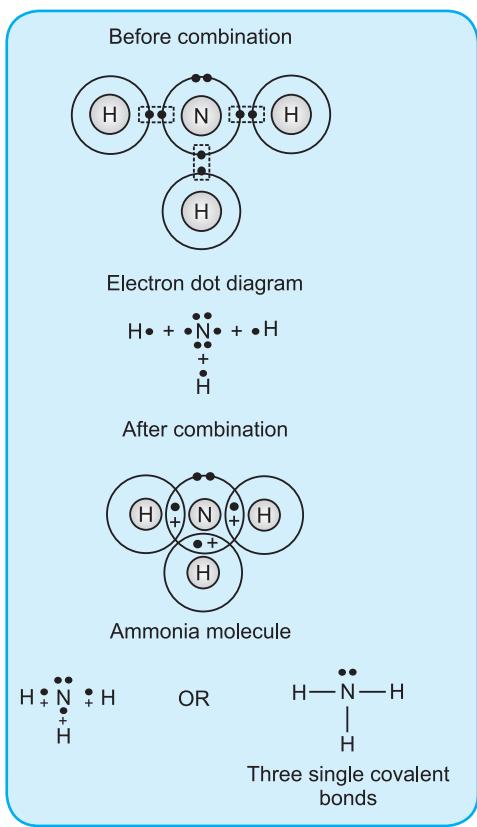


Fig. 2.9 : Formation of Ammonia Molecule

The hydrogen atom has 1 electron in its valence shell and it needs 1 electron to attain duplet configuration. Hence, the three hydrogen atoms share one electron with one nitrogen atom and the nitrogen atom shares 3 electrons with the three hydrogen

atoms to attain stable structure. This gives one lone pair of electrons on nitrogen and a single bond is formed between each of the 3 hydrogen atoms and the nitrogen atom.

#### Formation of Water Molecule ( $\text{H}_2\text{O}$ )

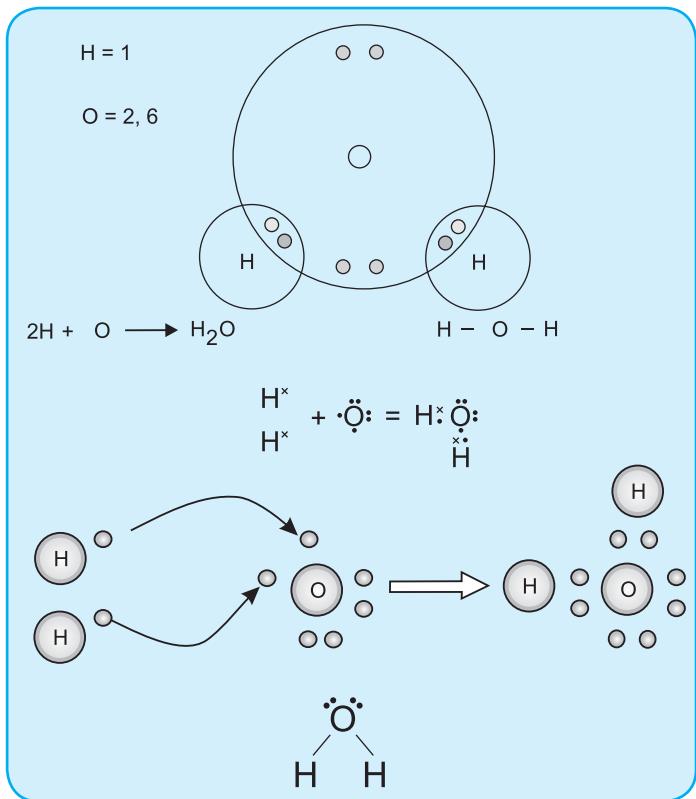


Fig. 2.10 : Formation of Water Molecule

A molecule of water is formed by the sharing of two electron pairs between two hydrogen atoms and one oxygen atom (see Figure 2.10). The oxygen atom has 6 electrons in its valence shell. So, it needs 2 electrons to attain stable octet structure. Each hydrogen atom has one electron in its valence shell, so it needs one electron to attain the stable duplet structure. Hence, the mutual sharing of electrons takes place between the atoms to form a stable structure.

#### Formation of methane molecule ( $\text{CH}_4$ )

Carbon has 4 valence electrons in its valence shell (see Figure 2.11). It needs 4 electrons to attain the stable electronic configuration of the nearest noble gas (for carbon, neon is the nearest noble gas). So, it shares 4 pairs of electrons with 4 hydrogen atoms (each of which has one electron in its valence shell, needing 1 electron to attain stable duplet structure).

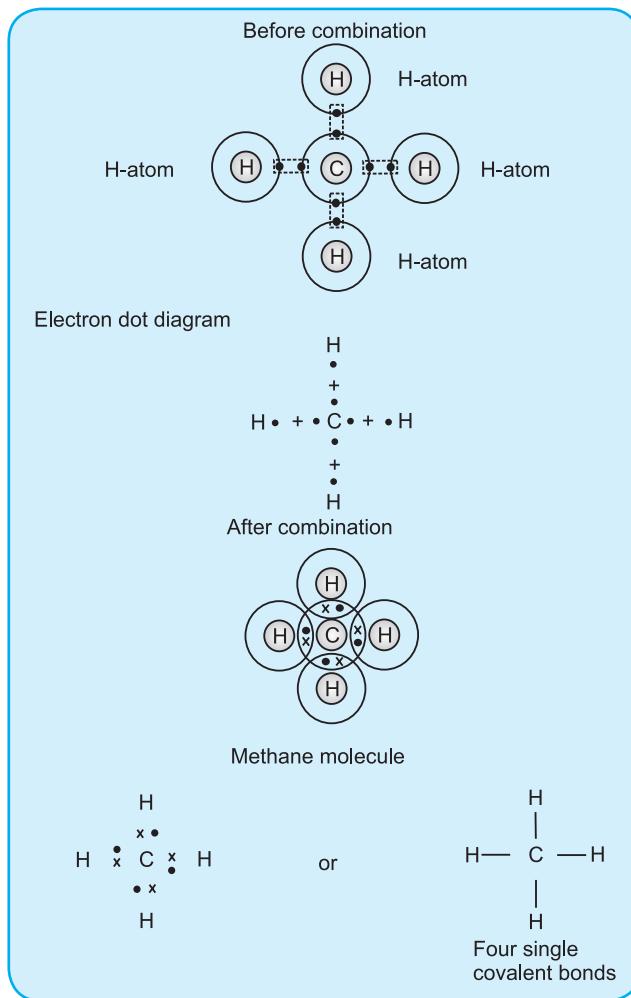
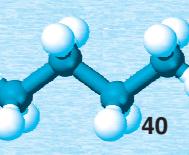


Fig. 2.11 : Formation of Methane Molecule

### Formation of Chlorine Molecule ( $\text{Cl}_2$ )

Chlorine atom has 7 electrons in the valence shell. It needs 1 electron to attain stable octet structure (see Figure 2.12). Since, we have two chlorine atoms, each atom shares 1 electron between themselves. Here the chlorine atoms share the unpaired electrons and attain the stable octet structure.

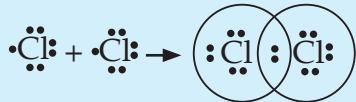


Fig. 2.12 : Formation of Chlorine Molecule

The valence electron which is not shared between the atoms is known as **non-bonding electrons** or in pairs, **lone pairs**. Here one covalent bond is formed between the atoms and 3 lone pairs of electrons are present on each of them.

### Formation of Carbon Tetrachloride Molecule ( $\text{CCl}_4$ )

In a carbon tetrachloride molecule a carbon atom

shares its 4 unpaired electrons with 4 chlorine atoms, each of which has 1 unpaired electron to share(see figure 2.13).

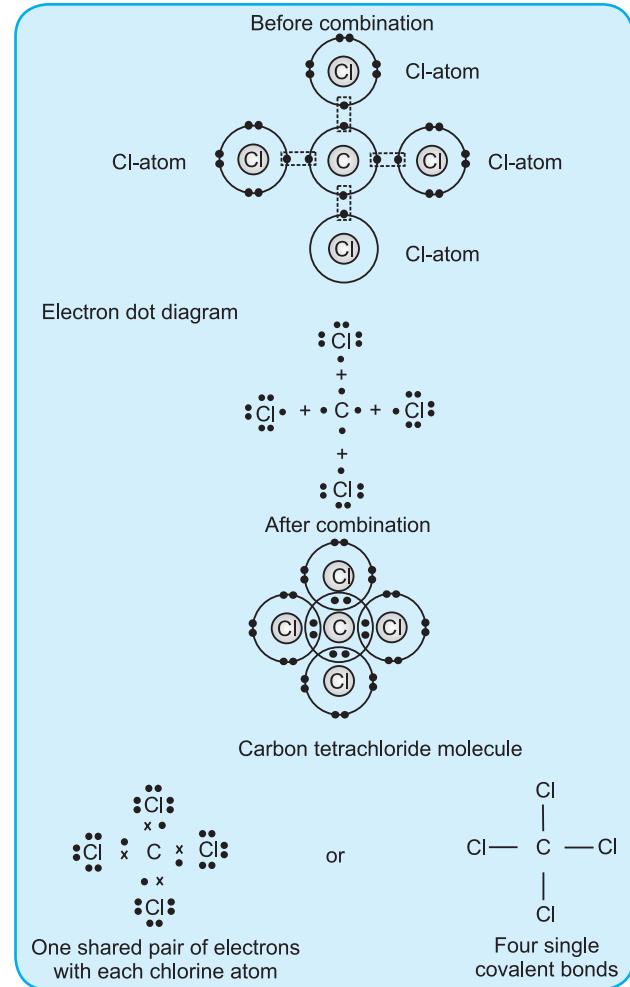
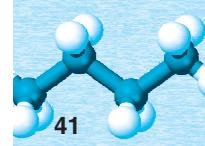


Fig. 2.13 : Formation of Carbon Tetrachloride Molecule

### Properties of covalent compounds

1. They are gases, liquids or soft solids. As only weak forces exist between the molecules of most covalent compounds, they exist as gases, liquids and soft solids.
2. Covalent compounds have low melting and boiling points. This is due to the weak intermolecular force between the atoms. A little energy is needed to overcome this force resulting in low melting points and boiling points in covalent compounds.
3. They are usually soluble in organic solvents like benzene, alcohol, acetone, chloroform and carbon tetrachloride rather than in water. This is because the attraction between the molecules of an organic solvent is comparable between the solvent molecules and covalent solute molecules (both behave as non-polar molecules, as most of the covalent compounds are non-polar).



This allows the solvent molecules to mix with the solute molecules. Covalent compounds are generally not soluble in water because there is no attraction between a covalent molecule and a water molecule (water is a polar solvent whereas, most of the covalent compounds are non-polar). However, some covalent compounds, especially the polar ones like ammonia and hydrochloric acid are highly soluble.

- They undergo reactions slowly. This is in contrast to the reactions of ionic compounds in solution, which are instantaneous. Reactions involving covalent compounds are slow as the breaking and formation of bonds in the reacting covalent molecules takes time.

## 2.5. DIFFERENCE BETWEEN IONIC AND COVALENT COMPOUNDS

**Table 2.7 Difference between Ionic and Covalent Compounds**

Ionic compound	Covalent compound
They are always solids.	They are generally gases and liquids.
They are formed by the complete transfer of electrons.	They are formed by mutual sharing of electrons.
They contain ions.	They do not contain ions.
They have high melting points and boiling points.	They have low melting points and boiling points.
In a solid state, they do not conduct electricity, but in the liquid state they conduct electricity.	They do not conduct electricity.
They are soluble in water but insoluble in organic solvents.	They are insoluble in water, but soluble in organic solvents.
They do not exhibit isomerism.	They exhibit isomerism.
Reactions between ionic compounds are rapid in solution.	Reactions between covalent compounds are generally slow.

## 2.6. REDOX REACTION

Chemical reactions involving simultaneous oxidation and reduction are called '**redox reactions**'.

The removal of one or more electrons from a particle (atom, molecule or ion) is called oxidation.

The addition of one or more electrons to a particle (atom, molecule or ion) is known as reduction.

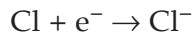
### For Example :

In the process of oxidation of sodium atom

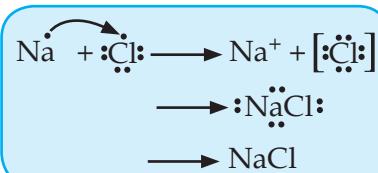


The sodium atom loses an electron to become  $\text{Na}^+$  ion. This reaction is an oxidation reaction.

In the process of reduction of chlorine atom

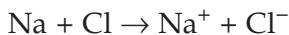


A chlorine atom gains an electron to form a chloride ion. This reaction is a reduction reaction (see Figure 2.14).



**Fig. 2.14 : Redox Reaction**

Thus the reaction



is a redox reaction.

An atom which gives one or more electrons to the other is called a reducing agent. In the above example Na is the reducing agent.

An atom which accepts one or more electrons from the other is called an oxidizing agent. In the above example, Cl is the oxidizing agent.

## 2.7. COORDINATE BOND

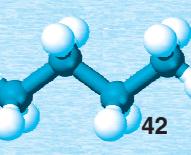
The electron contributions of combining atoms in a covalent bond are generally equal. In each shared pair of electrons, one electron is contributed from each atom of the bond. However, in some bond formation, the whole of the shared pair of electrons come from only one of the combining atoms of the bond, which is referred to as the donor atom. The other atom which does not contribute any electron to the shared pair, but tries to pull the pair of electron towards itself is called as the acceptor atom. This is called as the **coordinate bond or dative bond**.

The coordinate bonds are represented by an arrow symbol which points from the donor to the acceptor atom.

### Formation of $\text{H}_3\text{O}^+$ ion

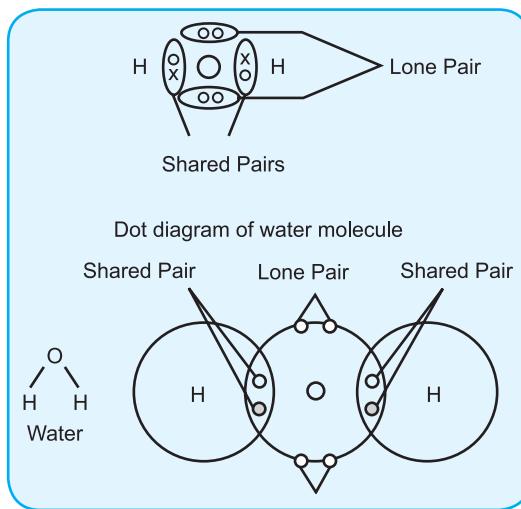
When an acid dissolves in water, an hydronium ( $\text{H}_3\text{O}^+$ ) ion is formed.

Water contains 2 hydrogen atoms and 1 oxygen atom. The hydrogen atom has 1 electron in its valence



shell while the oxygen atom has 6 electrons in its valence shell.

Two hydrogen atoms share 1 electron each with an oxygen atom to form a water molecule.



**Fig. 2.15 : Formation of  $\text{H}_3\text{O}^+$  Ion**

The two unshared pairs of electrons on oxygen, known as lone pair electrons, do not participate in the bond formation in the water molecule.

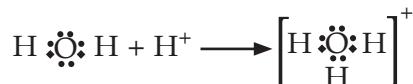
The OH bonds in a water molecule are polarised due to the large difference in the electronegativity values of oxygen and that of the hydrogen atoms. Oxygen has higher electronegativity, so it gains a partial negative charge while the hydrogen atoms attain partial positive charge. Therefore, a molecule of water behaves as a polar molecule because of these charges and is represented as shown in the Figure 2.15.

When an acid is added to water, strong dipole interaction (attraction between partial positive and negative charge) occurs between the polar water molecules and the polar acid molecules.

As a result of such a dipole interaction, the hydrogen ion gets released from the polar bonds of the acid molecules.



The hydrogen ion spontaneously adds onto one of the two lone pairs of electrons of the oxygen atom in the water molecule.

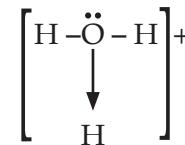


This results in the formation of hydronium ion by the addition of  $\text{H}^+$  ions to a lone pair of electrons of the oxygen atom of the polar water molecule.

Hydronium ion is hydrated proton:

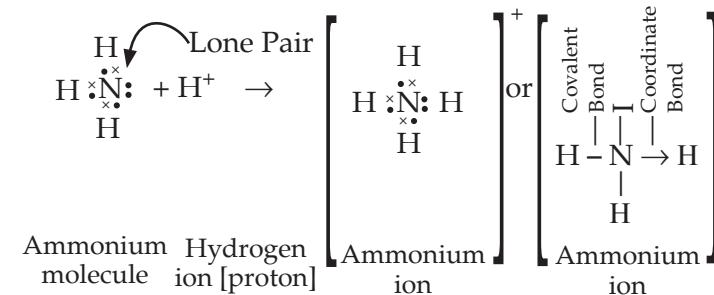


Structure of hydronium ion



### Formation of ammonium ion ( $\text{NH}_4^+$ )

Ammonia, on reacting with an acid or water, forms an  $\text{NH}_4^+$  ion.



**Fig. 2.16 : Formation of Ammonium Ions ( $\text{NH}_4^+$ )**

In this reaction, the lone pair of electron on N in  $\text{NH}_3$  is donated to the H of the acid or water and it attains duplet electronic configuration. The H atom leaves its own electron with Cl or OH. Chlorine and OH become negatively charged, and  $\text{H}^+$  attaches itself to  $\text{NH}_3$  to form the  $\text{NH}_4^+$ . (See Figure 2.16.)

The nitrogen atom in a molecule contains a lone pair of electrons, it is a DONOR. The hydrogen ion  $\text{H}^+$  accepts that lone pair of electrons (ACCEPTOR) and forms a coordinate bond. Once this bond is formed, all four-N—H bonds become identical. The linkage forms an ammonium ion having a single positive charge.

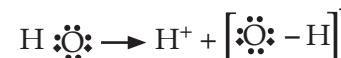
### Formation of hydroxyl ion ( $\text{OH}^-$ ) :

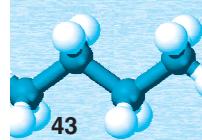
The hydroxyl ion is formed by removing one hydrogen ion ( $\text{H}^+$ ) from the water molecule.



When the  $\text{H}^+$  ion is removed from the water molecule, the shared pair of electron remains with oxygen, as oxygen is more electronegative. Thus, hydroxyl ion has a negative charge.

Structure of hydroxyl ion





## INTEXT QUESTIONS

1. Differentiate between atoms and ions.
2. Give the lewis dot structure of hydrogen, oxygen, chlorine
3. Give the structure of sodium chloride.
4. Give the diagrammatic view for the formation of :
  - (a) Oxygen molecule
  - (b) Nitrogen molecule
5. Give the properties of covalent compounds.
6. Differentiate between covalent and ionic compounds.
7. Give the concept behind the formation of a coordinate bond.



## POINTS TO REMEMBER

- ◆ The tendency of an element to achieve the nearest inert gas configuration is called the octet rule or octet law or the electronic theory of valency.
- ◆ Types of bonds :  
Ionic bonds, covalent bonds and coordinate bonds.
- ◆ Ionic bonds are formed between the atoms of an electropositive element and an electronegative element. One or more electrons are completely transferred from the electropositive atom to the electronegative atom.
- ◆ Covalent bonds are formed by the mutual sharing of electrons between the atoms.
- ◆ Ionic compounds are soluble in water and they conduct electricity in aqueous form, but not in solid form as there is no mobility of electrons in solid state.
- ◆ Each shared pair of electrons is equivalent to one covalent bond.
- ◆ The electron pair which is not taking part in a reaction is called the lone pair of electrons.
- ◆ Covalent compounds are generally liquids, gases or low-melting solids because of the weak intermolecular force between the atoms.
- ◆ A maximum of three covalent bonds can be formed between two atoms.

## EXERCISES

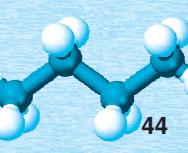
### A. FILL IN THE BLANKS

1. The number of electrons lost, gained or shared by an atom of an element is known as it's .....
2. The electrostatic forces of attraction between oppositely charged ions is known as .....
3. Ionic compounds have ..... volatility.
4. Ionic compounds are generally ..... in polar solvents like water.
5. An atom of Ca loses two electrons. It attains the structure of .....
6. All the inert gases except helium have ..... electrons in their outermost orbits.
7. The melting and boiling points of covalent compounds are generally .....
8. In the formation of an ammonia ion, ..... is the donor.
9. ..... is an example of a triple covalent bond.
10. The electric bulb is filled with the gas .....

### B. MULTIPLE CHOICE TYPE QUESTIONS

**Choose the Correct Answer from the options given below :**

1. Bond formation occurs when the energy of the system.....  
(a) Decreases                                  (b) Increases  
(c) Remains the same
2. The element that obeys duplet law is :  
(a) O    (b) H  
(c) N    (d) C
3. The number of electrons in chloride ion is :  
(a) 16    (b) 17  
(c) 18
4. A double bond contains  
(a) Two shared electrons  
(b) Four shared electrons  
(c) Three shared electrons



5. Which one of the following is bad conductor of electricity?
- Ionic compound
  - Covalent compound
  - Coordinate compound

#### C. ANSWER THE FOLLOWING QUESTIONS BASED ON THE INSTRUCTIONS GIVEN BELOW

- Write the electron dot structures of the following :
  - Hydrogen
  - Oxygen
  - Chlorine
- Differentiate between sodium atom and ion on the basis of colour and toxicity.
- Draw the structure for the formation of hydrogen molecule.
- List two points of difference between covalent and ionic compounds.

#### D. MATCH THE COLUMNS

Column I	Column II
1. Electrovalent bonds	Solids
2. Sodium atom	One pair of electrons
3. Covalent bonds	Mutual sharing of electrons
4. Ammonia molecule	Strong
5. Ionic bonds	Silvery white

#### E. ANSWER THE FOLLOWING QUESTIONS BASED ON THE OBSERVATIONS

- Covalent compounds are formed by the mutual sharing of electrons comment on it.
- Coordinate bond formation depends upon the difference between the electronegativities of the two atoms, explain with an example.

#### F. VERY SHORT QUESTIONS

- Name the charged particles which attract one another to form electrovalent compounds.
- The electronic configuration of nitrogen is 2, 5. How many electrons in the outer shell of a nitrogen atom are not involved in the formation of a nitrogen molecule ?
- Write the formula of the sulphate of the element with atomic number 13 ?
- What is a lone pair of electrons ?
- Name a neutral covalent molecule which contains one lone pair of electrons.

- Which covalent compound behaves like an ionic compound in aqueous solution ?
- What type of bonding is present in the water molecule ?
- Which covalent compound behaves like an ionic compound in aqueous solution ?

#### G. SHORT QUESTIONS

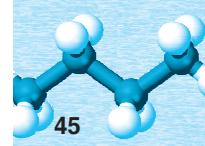
- Define a chemical bond.
- Define ionic or electrovalent bond.
- Define covalent bond.
- Define coordinate bond.
- How is an electrovalent bond formed ?
- Draw the Lewis dot structure for the following covalent bond formation
 

(a) Oxygen	(b) Methane
------------	-------------
- What is a single bond ?
- What is a double bond ?
- What is a triple bond ?
- Why ionic solids are non-conductors of electricity ?
- Why covalent solids are conductors of electricity ?

#### H. LONG QUESTIONS

- Explain the formation of ionic bonds in sodium chloride with the lattice structure.
- Write the difference between ionic and covalent compounds ?
- What are the properties of a covalent compound ?
- Give the electron dot and molecular structure of :
 

(a) Magnesium chloride	(b) Nitrogen
------------------------	--------------
- Predict the type of bonding in the following molecules :
   
Oxygen, Calcium Oxide, Water, Methane, Ammonium ion.
- Acids dissolve in water and produce positively charged ions. Draw the structure of these positive ions.
- Explain why  $\text{CCl}_4$  does not dissolve in water.
- Explain the bonding in methane molecule using electron dot structure.



## RECENT BOARD QUESTIONS

### Question 1.

**(a) Fill in the blanks :**

- (i) In covalent compounds, the bond is formed due to the ..... of electrons. [2011]
- (ii) Electrovalent compounds have a..... boiling point. [2011]
- (iii) A molecule of.....contains a triple bond. [2011]
- (iv) Electrovalent compounds have \_\_\_\_\_ melting points. [2016]

**(b) Choose the correct answer from the options given below :**

- (i) Identify the weak electrolyte from the following :  
[2012]
  - (A) Sodium Chloride solution
  - (B) Dilute Hydrochloric acid
  - (C) Dilute Sulphuric acid
  - (D) Aqueous acetic acid
- (ii) Among the following compounds identify the compound that has all three bonds (ionic, covalent and coordinate bond).  
[2013]
  - (A) Ammonia
  - (B) Ammonium chloride
  - (C) Sodium hydroxide
  - (D) Calcium chloride
- (iii) Which of the following is not a typical property of an ionic compound?  
[2013]
  - (A) High melting point
  - (B) Conducts electricity in the molten and in the aqueous solution state.
  - (C) They are insoluble in water
  - (D) They exist as oppositely charged ions even in the solid state.
- (iv) Bonding in this molecule can be understood to involve coordinate bonding :  
[2015]
  - (A) Carbon tetrachloride
  - (B) Hydrogen
  - (C) Hydrogen chloride
  - (D) Ammonium chloride
- (v) The particles present in strong electrolytes are :  
[2016]
  - (A) only molecules
  - (B) mainly ions
  - (C) ions and molecules
  - (D) only atoms

**(c) Write balanced chemical equations for the following :**

- (i) A mixture of sodalime and sodium acetate is heated  
[2011]
- (ii) Calcium carbide to ethyne  
[2015]
- (iii)  $\text{Fe} \xrightarrow{\text{A}} \text{FeCl}_3 \xrightarrow{\text{B}} \text{FeCO}_3 \xrightarrow{\text{C}} \text{Fe}(\text{NO}_3)_2$   
[2015]
- (iv) Sodium ethanoate to methane  
[2015]

### Question 2.

- (a) (i) Differentiate between electrical conductivity of copper sulphate solution and copper metal.  
[2011]

(ii) By drawing an electron dot diagram, show the lone pair effect leading to the formation of ammonium ion from ammonia gas and hydrogen ion.  
[2011]

**(b)** Name the method used for preparation of the following salts from the list given below : [2011]

- (i) Sodium nitrate
- (ii) Iron (III) chloride
- (iii) Lead chloride
- (iv) Zinc sulphate
- (v) Sodium hydrogen sulphate

List :

- (A) Simple displacement
- (B) Neutralization
- (C) Decomposition by acid
- (D) Double decomposition
- (E) Direct synthesis.

### Question 3.

**(a) Name the gas in each of the following :** [2012]

- (i) The gas evolved on reaction of Aluminium with boiling concentrated caustic alkali solution.
- (ii) The gas produced when excess ammonia reacts with chlorine.
- (iii) A gas which turns acidified potassium dichromate clear green.
- (iv) The gas produced when copper reacts with concentrated nitric acid.
- (v) The gas produced on reaction of dilute sulphuric acid with a metallic sulphide.

**(b) State one observation for each of the following :** [2012]

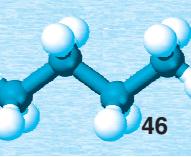
- (i) Excess ammonium hydroxide solution is added to lead nitrate solution.
- (ii) Bromine vapours are passed into a solution of ethyne in carbon tetrachloride.
- (iii) A zinc granule is added to copper sulphate solution.
- (iv) Zinc nitrate crystals are strongly heated.
- (v) Sodium hydroxide solution is added to ferric chloride solution at first a little and then in excess.

### Question 4.

- (a) (i) Why do covalent compounds exist as gases, liquids or soft solids ?  
[2014]
- (ii) Which electrode : anode or cathode is the oxidising electrode ? Why?  
[2014]

**(b)** M is a metal above hydrogen in the activity series and its oxide has the formula  $\text{M}_2\text{O}$ . This oxide when dissolved in water forms the corresponding hydroxide which is a good conductor of electricity. In the above context answer the following :  
[2014]

- (i) What kind of combination exists between M and O ?



- (ii) How many electrons are there in the outermost shell of M ?
- (iii) Name the group to which M belongs.
- (iv) State the reaction taking place at the cathode.
- (v) Name the product at the anode.

(c) **Answer the following questions :**

- (i) Give suitable chemical terms for the following
  - (a) A bond formed by a shared pair of electrons with both electrons coming from the same atom. [2013]
  - (b) When crystals of copper nitrate are heated in a test tube. [2015]

**Question 5.**

(a) **An element L consists of molecules :** [2015]

- (i) What type of bonding is present in the particles that make up L ?

- (ii) When L is heated with iron metal, it forms a compound  $\text{FeL}$ . What chemical term would you use to describe the change undergone by L ?

(b) **The following table shows the electronic configuration of the elements W, X, Y, Z :** [2016]

Element	W	X	Y	Z
Electronic configurations	2, 8, 1	2, 8, 7	2, 5	1

Answer the following questions based on the table above :

- (i) What type of Bond is formed between :
  - (a) W and X
  - (b) Y and Z
- (ii) What is the formula of the compound formed between :
  - (a) X and Z
  - (b) W and X
- (iii) Give reasons why copper is a good conductor of electricity; although it is a non-electrolyte.



# 3

# STUDY OF ACIDS, BASES AND SALTS

## LEARNING OUTCOMES :

- 3.0. Introduction
- 3.1. Classification of Acids
- 3.2. Strength of Acids
- 3.3. Concentration of Acids
- 3.4. Hydracids and Oxyacids
- 3.5. Basicity of an Acid
- 3.6. Preparation of Acids
- 3.7. Properties of an Acid
- 3.8. Uses of Acids
- 3.9. Bases
- 3.10. The Acidity of a Base
- 3.11. Preparation of Bases
- 3.12. Properties of Bases
- 3.13. Uses of Bases

- 3.14. Ionic Theory of Acids, Bases and Salts
- 3.15. Neutralization of Acids and Bases
- 3.16. Practical Applications of Neutralization Reaction
- 3.17. Salts
- 3.18. Acid–Base Indicators
- 3.19. Universal Indicators and pH Paper
- 3.20. Acid Rain
- 3.21. General Method for Preparing Some Salt
- 3.22. Decomposition of Bicarbonates, Carbonates, Chlorides and Nitrates by Acids
- 3.23. Water of Crystallization
- 3.24. Efflorescence
- 3.25. Hygroscopic Substances
- 3.26. Deliquescence
- 3.27. Drying and Dehydrating Agents

## 3.0 INTRODUCTION

Have you wondered why the taste of vinegar or lemon is sour? This taste is due to chemical compounds present in them, called acids. In general, chemical compounds can be classified into three categories namely : Acids, Bases and Salts.

Acids and bases are commonly used in laboratories and industries. Salts are used both in industries and at home. These chemical compounds are useful in the preparation of some useful chemicals such as soaps, cosmetics, etc. In this chapter, we will study the properties and uses of acids, bases and salts.

### Acids

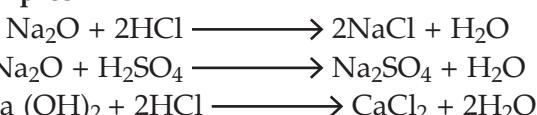
The word acid is derived from the Latin word 'acidus', which means "sour". Acids are chemical substances that contain replaceable hydrogen atoms. These replaceable hydrogen atoms can be replaced with atoms of a metal during neutralization with a base.

### Bases

A base can be defined as a substance that reacts with an acid to form salt and water.

Generally, bases are oxides or hydroxides ( $\text{OH}^-$ ) of metals.



**Examples :****Salts**

A salt is a compound formed by the replacement of the hydrogen of an acid by metals or by positive radicals, which comes from a base or an alkali. For example : the salt NaCl is formed by the replacement of the H atom of HCl by Na (metal) atom of NaOH.

**3.1. CLASSIFICATION OF ACIDS**

Acids can be classified into two types based on their sources

- ◆ Naturally Occurring Acids or Organic Acids (Biological Origin)
- ◆ Mineral Acids or Inorganic Acids

**Organic Acids**

Acids which are obtained from animal and plant sources are called organic acids. They contain carbon atoms in them.

**Examples :**

Oxalic acid ( $\text{COOH}$ )<sub>2</sub>

Acetic acid ( $\text{CH}_3\text{COOH}$ )

Some common naturally occurring organic acids are listed below in the table :

Source	Acid
Grapes, lemons, oranges	Citric acid
Apples	Malic acid
Sting of ants and bees	Formic acid
Sour milk and yoghurt	Lactic acid
Vinegar	Acetic acid
Proteins	Amino acid
Tea	Tannic acid
Tamarind and gooseberries	Tartaric acid
Tomatoes	Oxalic acid
Fats	Stearic acid
Lettuce, vegetable and unripe fruits	Succinic acid
Olive oil	Oleic acid
Urine	Uric acid

These acids are also considered as weak acids because they don't completely ionize in the solution. So their solutions contain both ions and molecules in

them.

**Inorganic Acids :**

These are the acids which are obtained from minerals. Hence, they are also called "mineral acids".

Based on their molecular composition, inorganic acids are classified into hydracids or oxyacids. Hydracids do not contain oxygen and are binary compounds with only two elements in them, e.g., HCl. Oxyacids contain oxygen, hydrogen and another element in them.

Inorganic acids in general do not contain carbon except carbonic acid ( $\text{H}_2\text{CO}_3$ ), which is present in soft drinks and is non-corrosive in nature (weak acid).

**Examples :**

Hydrochloric acid (HCl)

Sulphuric acid ( $\text{H}_2\text{SO}_4$ )

Nitric acid ( $\text{HNO}_3$ )

They are strong acids as they ionize completely in solution.

**3.2. STRENGTH OF ACIDS**

The strength of an acid depends upon the degree of ionization [ $\alpha$ ] and concentration of hydronium ions  $[\text{H}_3\text{O}]^+$  produced by the acid in aqueous solution.

**Strong Acids**

They completely ionize in aqueous solution, thereby producing higher concentration of hydronium ions  $[\text{H}_3\text{O}]^+$ .

**Examples :**

Hydrochloric acid (HCl)

Sulphuric acid ( $\text{H}_2\text{SO}_4$ )

Nitric acid ( $\text{HNO}_3$ )

**Weak Acids**

They ionize partially in aqueous solutions, resulting in the formation of ions as well as molecules in solution. They produce low concentration of hydronium ions  $[\text{H}_3\text{O}]^+$  in aqueous solution.

**Examples :**

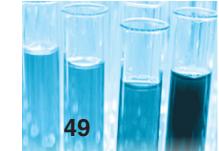
Acetic acid, carbonic acid, etc.,

**Note :**

- ◆ Degree of ionization

$$\frac{\text{No. of acid molecules ionised}}{\text{Total no. of acid molecules present in aqueous solution}} \times 100$$

- ◆ If the degree of ionization for an acid, base or salt in aqueous solution is greater than 30%, it is considered as a strong acid and if it is less than 30%, it is known as a weak acid.



### 3.3. CONCENTRATION OF ACIDS

Acids are classified into two types, depending on their concentration, as dilute acids and concentrated acids.

#### Dilute Acids

Dilute acids contain more water than the amount of acid present in them.

#### Concentrated Acids

Concentrated acids contain very little amount of water or no water at all.

#### NOTE

*The concentration of an acid simply tells us the amount of water present in the acid and not the strength of the acid. The strength of an acid is the measure of concentration of hydronium ions it produces in its aqueous solution. Thus, dilute HCl is stronger than highly concentrated acetic acid.*

#### Procedure For Dilution of an acid

The concentrated acid can be diluted by pouring the concentrated acid little by little into water with constant stirring, using a glass rod. The reverse, i.e., water is not to be added to the acid, because often the reaction is highly exothermic. Therefore, a lot of heat energy is produced and splashing of acid may occur.

### 3.4 HYDRACIDS AND OXYACIDS

Acids can be classified based upon their molecular composition.

#### Hydracids

Hydracids are acids which consist of hydrogen and a non-metallic element other than oxygen in their molecules.

Examples for hydracids are HCl, H<sub>2</sub>S, HI, HBr and HF.

#### Oxyacids

Oxyacids are acids which contain oxygen along with hydrogen and another element.

Examples of oxyacids are

HNO<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub>, H<sub>3</sub>PO<sub>4</sub>, H<sub>2</sub>CO<sub>3</sub>, H<sub>2</sub>SO<sub>3</sub>.

### 3.5 BASICITY OF AN ACID

The basicity of an acid is the number of hydrogen ions [H<sup>+</sup>] which can be produced per molecule of the acid in aqueous solution or the number of hydroxyl ions (of a base) with which one molecule of an acid can combine.

Based upon the basicity, acids can be classified as monobasic, dibasic, tribasic, and so on.

#### Monobasic Acids :

Acids, which produce one hydronium ion per molecule of the acid, by ionization with water, are known as monobasic. Some examples of monobasic acids are HCl, HNO<sub>3</sub>, HBr, HI, HClO, HClO<sub>3</sub>, HClO<sub>4</sub>, HCOOH and CH<sub>3</sub>COOH.

Examples :

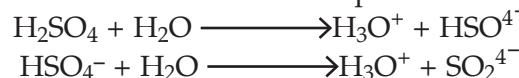


Here for an acetic acid molecule there are four hydrogen atoms, but it is a monobasic acid. The reason for this is that the molecule upon ionization liberates only one hydronium ion.

#### Dibasic Acids :

Acids, which produce two hydronium ions (H<sub>3</sub>O<sup>+</sup>) per molecule of the acid, by ionization with water, are known as dibasic acids. Some examples of dibasic acids are H<sub>2</sub>SO<sub>4</sub>, H<sub>2</sub>C<sub>2</sub>O<sub>4</sub>, H<sub>2</sub>SO<sub>3</sub>, H<sub>2</sub>CO<sub>3</sub>.

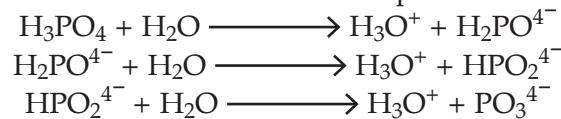
Dibasic acids ionize in two steps



#### Tribasic Acids :

Acids that produce three hydronium ions per molecule of the acid on ionization with water are known as tribasic acids. Some examples of tribasic acids are H<sub>3</sub>PO<sub>4</sub>, H<sub>3</sub>PO<sub>3</sub>.

Tribasic acids ionize in three steps :



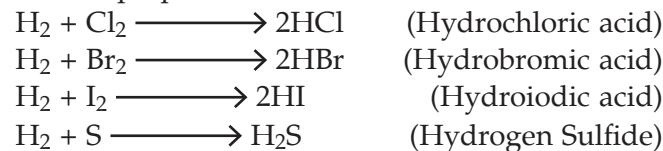
### 3.6. PREPARATION OF ACIDS

Acids can be prepared by the following general methods :

1. By synthesis or direct combination
2. By the action of water on non-metallic oxides
3. By displacement
4. By oxidation
5. By hydrolysis

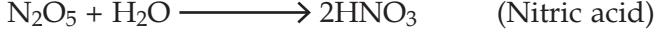
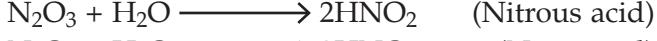
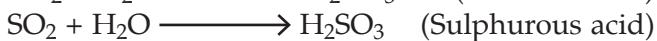
#### 1. Preparation of Acids (Synthesis or Direct Combination Method)

By the direct combination method, binary acids, i.e., acids consisting of two elements such as HCl, HBr can be prepared.



## 2. Preparation of Acids (Action of Water on Non-Metallic or Acidic oxides)

Acids can be prepared by the reaction of non-metals with water.



### NOTE

Only oxyacids can be prepared by this method. These acidic oxides are known as acidic anhydrides.

## 3. Preparation of Acids By Displacement Method

In this method an element reacts with a compound, displacing another element from it resulting in the formation of salt and acid.



## 4. Preparation of Acids By Oxidation Method

A non-metal like sulphur or phosphorous is oxidized by concentrated nitric acid to form sulphuric acid or phosphoric acid



## 3.7. PROPERTIES OF AN ACID

- Most of the acids are generally soluble in water.
- Some acids are solids and some are liquids at room temperature.

**For example :**

Acid	Formula	Physical State
Boric acid	H <sub>3</sub> BO <sub>3</sub>	Solid Acids
Citric acid	C <sub>6</sub> H <sub>8</sub> O <sub>6</sub>	
Oxalic acid	(COOH) <sub>2</sub>	
Phosphoric acid	H <sub>3</sub> PO <sub>4</sub>	
Tartaric acid	C <sub>4</sub> H <sub>6</sub> O <sub>6</sub>	
Acetic acid	CH <sub>3</sub> COOH	Volatile acids
Formic acid	HCOOH	vaporize easily at room temperature
Carbonic acid	H <sub>2</sub> CO <sub>3</sub>	or on heating at 100°C
Hydrochloric acid	HCl	
Nitric acid	HNO <sub>3</sub>	Liquid Acids
Sulphurous acid	H <sub>2</sub> SO <sub>3</sub>	
Sulphuric acid	H <sub>2</sub> SO <sub>4</sub>	Non-volatile Liquid Acids

3. All mineral acids, for example, sulphuric acid and nitric acid are corrosive in nature whereas weak acids like carbonic acid and organic acids are not corrosive.

4. They turn blue litmus into red.

5. Their aqueous solutions conduct electricity as, they are electrolytes.

6. They react with bases to form salt and water.

**Example :**



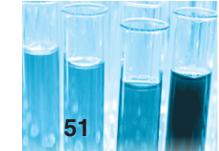
7. Carbonate and hydrogen carbonates produce CO<sub>2</sub> gas with effervescence upon reaction with acids.



## 3.8. USES

### Acids

Acids	Uses
Hydrochloric acid (HCl)	It is used to prepare chlorides and for pickling (removing oxide from the surface of iron and steel before plating) Diluted form is used to clean sanitary ware Used as a bleaching agent
Sulphuric acid (H <sub>2</sub> SO <sub>4</sub> )	Used in car batteries used in manufacture of detergents, drugs, dyes and paints Manufacture of fertilizers such as ammonium sulphate and super phosphate
Nitric acid (HNO <sub>3</sub> )	Production of fertilizers such as ammonium nitrate Manufacture of explosives such as TNT and nitroglycerine purification of gold and silver ornaments
Carbonic acid	In soft drinks to give fizz
Citric acid	Source of vitamin C especially in medicines Also in food preservation and in flavoring soft drinks
Boric acid	Antiseptic and in grain preservation
Acetic acid	Cooking, flavoring and preserving food
Oxalic acid	Constituent of stain remover
Tartaric acid	Constituent of baking powder



Benzoic acid	Food preservation, manufacture of perfumes and medicines
Phosphoric acid	Constituent of fertilizers

*Presence of H<sub>2</sub> in acid can be identified by adding Mg ribbon that will liberate H<sub>2</sub> gas to know whether a substance is an acid, just add any form of carbonate that will result in the liberation of CO<sub>2</sub>.*

### 3.9. BASES

A base can be defined as a substance that reacts with an acid to form salt and water. A base that is soluble in water is known as an alkali. Alkalies dissolve in water to give hydroxyl ions as the only negative ions. Examples : NaOH, KOH.

**All alkalis are bases but all bases are not alkalis**

In general, bases are oxides or hydroxides of metals. And they also combine with the hydronium ion of acids to form salt and water. Some examples of common bases are sodium hydroxide, potassium hydroxide, calcium oxide, calcium hydroxide and aluminum hydroxide.

The bases are classified based on their strength. They are strong bases and weak bases.

#### Strong Bases

Bases which dissociate almost completely to give OH<sup>-</sup> ions in aqueous solutions are called strong bases.

For example : NaOH, KOH and Be(OH)<sub>2</sub> are strong bases as they almost get completely dissociated.

#### Weak Bases

Bases which dissociate only to a small extent to give OH<sup>-</sup> ions in aqueous solutions are known as weak bases.

#### For example :

NH<sub>3</sub>, NH<sub>4</sub>OH and sparingly soluble metal oxides and hydroxides; ZnO and Al(OH)<sub>3</sub> are weak bases as they dissociate only to a small extent to give OH<sup>-</sup> ions.



### 3.10. THE ACIDITY OF A BASE

The acidity of a base is determined as the number of hydroxyl ions that one molecule of that base gives on dissociation in water.

OR

Acidity of a base may be defined as the number of replaceable hydroxyl groups per molecule of the base. There are three types of bases

- ◆ Monoacidic base
- ◆ Diacidic base
- ◆ Triacidic base

#### Monoacidic Base

Monoacidic base contains one replaceable hydroxyl ion (OH<sup>-</sup>) per molecule of the water-soluble base.

#### Example :



#### Diacidic Base

Diacidic base contains two replaceable hydroxyl (OH<sup>-</sup>) ions per molecule of the water-soluble base.

#### Example :



#### Triacidic Base

Triacidic base contains three replaceable hydroxyl ions (OH<sup>-</sup>) per molecule of the water-soluble base.

#### Example :

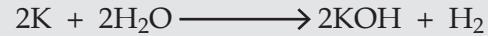
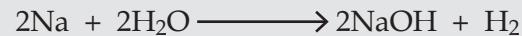


### 3.11. PREPARATION OF BASES

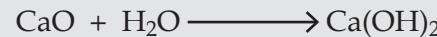
#### 1. Metals react with oxygen to give bases.



#### 2. Highly reactive metals like sodium, potassium and calcium react with water to form bases.



#### 3. Soluble metallic oxides like sodium oxide, potassium oxide and calcium oxide react with water to form bases.



#### 4. Decomposition of Salts : On decomposition of metallic carbonates, basic oxides and carbon dioxide are formed.

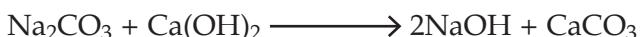
Metal carbonate	$\xrightarrow{\Delta}$	Basic oxide + Carbon dioxide
$\text{CaCO}_3$	$\longrightarrow$	$\text{CaO} + \text{CO}_2$
$\text{CuCO}_3$	$\longrightarrow$	$\text{CuO} + \text{CO}_2$

5. Double Decomposition : Metallic hydroxide is obtained with salt when aqueous solutions of salts react with a base.

#### Salt solution + Base $\longrightarrow$ Hydroxide + Salt



6. Soluble sodium hydroxide is prepared by heating sodium carbonate with slaked lime.



7. Ammonium hydroxide can be prepared by simply dissolving ammonia gas in water. Ammonia is extremely soluble in water.



### 3.12 PROPERTIES OF BASES

1. Bases are generally colourless and odourless. However, ammonium hydroxide solution has a pungent odour.

2. Bases like sodium hydroxide and potassium hydroxide dissolve in water producing heat. Hence, these reactions are called as exothermic reactions.

3. Aqueous solutions of bases are soapy to touch and are corrosive in nature.

4. They have a bitter taste.

5. They turn red litmus into blue, methyl orange to yellow and colourless phenolphthalein to pink.

6. In aqueous state they conduct electricity, and so are called as strong electrolytes.

7. Bases react with ammonium salts and liberate ammonia gas.



### 3.13 USES OF BASES

Base	Uses
Sodium hydroxide	Manufacture of soap
Calcium oxide (quick lime)	Neutralization of acidic nature of the soil caused by acid rain
Calcium hydroxide (slaked lime)	Preparing mortar and for white washing, manufacturing of bleaching powder, etc., Softening of hard water

Magnesium hydroxide	For neutralizing acidity produced by HCl present in the stomach (antacid)
Ammonium hydroxide	Removing the stains like grease, color, etc., from the clothes
Lithium hydroxide	Helps Astronauts in spaceships to neutralize dangerous levels of CO <sub>2</sub>
Potassium hydroxide	Added in alkaline batteries

The growth of crops gets hampered in acidic soil and slightly alkaline soil is best suited for plant growth. Thus, farmers use CaOH and powdered lime to make acidic soil alkaline as lime being a base reacts with acid to produce salt and water and the soil gets neutralized. It also helps in improving the texture of heavy soils.

### 3.14. IONIC THEORY OF ACIDS, BASES AND SALTS

The Swedish chemist, Svante Arrhenius framed the concept of acids and bases. He defined acids and bases in terms of the effect these substances have on water. According to Arrhenius, acids are substances that increase the concentration of H<sup>+</sup> ions in aqueous solution, and bases increase the concentration of OH ions in aqueous solution.

An acid is a substance that when dissolved in water, increases the concentration of hydronium ion, H<sub>3</sub>O<sup>+</sup> (aqueous). But chemists used H<sup>+</sup> (aqueous) instead of H<sub>3</sub>O<sup>+</sup> (aq) and called it as the hydrogen ion. However, as the aqueous hydrogen ion is not a bare proton in water, but a proton chemically bonded to water, it is called the hydronium ion.

Example :



#### Strong and Weak Acids

##### Strong Acids

Acids which dissociate practically completely into H<sub>3</sub>O<sup>+</sup> ions in aqueous solutions are called strong acids. For example, mineral acids are strong acids as they dissociate almost completely into H<sub>3</sub>O<sup>+</sup> ions in aqueous solutions.

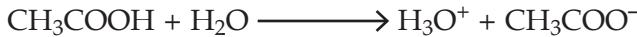
For example :



### Weak Acids

Acids which dissociate only to a small extent into  $\text{H}_3\text{O}^+$  ions in aqueous solutions are known as weak acids. For example, organic acids in general and many inorganic acids, like HF and  $\text{H}_2\text{CO}_3$ , dissociate only partially and to a small extent into  $\text{H}_3\text{O}^+$  ions.

#### For example :



### Bases

A Base is a substance that, when dissolved in water, increases the concentration of hydroxide ion,  $\text{OH}^-$  ion.

#### For example :



### 3.15. NEUTRALIZATION OF ACIDS AND BASES

An acid and a base neutralize each other to form salt and water.



Irrespective of the type of the acid or base, the neutralization reaction remains the same, i.e., the combination of  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$  to form  $\text{H}_2\text{O}$ .

### 3.16. PRACTICAL APPLICATIONS OF NEUTRALIZATION REACTION

- ◆ The acidity of the soil due to acid rain is reduced by adding slaked lime.
- ◆ Persons suffering from acidity are given antacid tablets, which contain magnesium hydroxide that neutralizes the acidity.
- ◆ When someone is stung by ants or bees (the sting injects formic acid into the body), the affected area can be neutralized by washing the part with soap, which is a base containing sodium hydroxide.

*An alkali can be identified by adding  $\text{NH}_4\text{Cl}$  to it and heating it. Ammonia gas is liberated to show the presence of alkali.*

Crops and plants grow better in a particular pH range; for example, rice grows better in slightly acidic soil, sugarcane in neutral and citrus fruits in alkaline soil.

### 3.17. SALTS

A Salt is a compound formed by replacing the hydrogen of an acid by a metal.

#### Example :



#### Types of Salts

There are six types of salts. They are as follows :

- (i) Normal salts
- (ii) Acidic salts
- (iii) Basic salts
- (iv) Mixed salts
- (v) Double salts
- (vi) Complex salts

#### 1. Normal Salts

They are formed by the complete replacement of the ionizable hydrogen ions of an acid by metal ions.

Some examples of normal salts and acids from which they are derived from are as follows :

Acids	Normal salts
$\text{HCl}$	$\text{NaCl}, \text{NH}_4\text{Cl}$
$\text{H}_2\text{SO}_4$	$\text{Na}_2\text{SO}_4, (\text{NH}_4)_2\text{SO}_4$
$\text{H}_3\text{PO}_4$	$\text{K}_3\text{PO}_4, (\text{NH}_4)_3\text{PO}_4$

#### 2. Acidic Salts

Salts formed by the partial replacement of the ionizable hydrogen ions of a polybasic acid by metal ions are known as acid salts.

Some examples of acid salts are as follows :

Acids	Acid salts
$\text{H}_2\text{SO}_4$	$\text{NaHSO}_4, \text{NH}_4\text{HSO}_4$
$\text{H}_2\text{CO}_3$	$\text{KHCO}_3, \text{Ca}(\text{HCO}_3)_2, \text{NH}_4\text{HCO}_3$
$\text{H}_3\text{PO}_4$	$\text{NaH}_2\text{PO}_4, (\text{NH}_4)\text{H}_2\text{PO}_4, \text{Na}_2\text{HPO}_4$

#### 3. Basic Salts

They are formed by the partial replacement of the ionisable hydrogen ions or the incomplete neutralization of a base.

Bases	Basic salts
$\text{Pb}(\text{OH})_2$	$\text{Pb}(\text{NO}_3)_2 \cdot \text{Pb}(\text{OH})_2$
$\text{Pb}(\text{OH})_2$	$\text{PbCO}_3 \cdot \text{Pb}(\text{OH})_2$
$\text{Cu}(\text{OH})_2$	$\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$
$\text{Mg}(\text{OH})_2$	$\text{MgCO}_3 \cdot \text{Mg}(\text{OH})_2 \cdot 3\text{H}_2\text{O}$

#### 4. Mixed Salts

They are formed by the combination of more than one acid or basic radical or more than one cation or anion.

**Examples :**  $\text{Ca}(\text{OCl})\text{Cl}$  – bleaching powder (calcium hypochlorite) it contains the anions  $\text{OCl}^-$  and  $\text{Cl}^-$ .

$\text{Na}_2\text{KPO}_4$ –Disodium potassium phosphate.

## 5. Double Salts

They are formed by the combination of two simple salts which dissolve in water to form a crystal.

Examples :



## 6. Complex Salts

They are formed when anions, neutral molecules, or both are, held together by a metal ion.

Examples :  $\text{K}_4\text{Fe}(\text{CN})_6$  – Potassium ferrocyanide.

$\text{NaAlO}_2$  – Sodium aluminate

## 3.18. ACID-BASE INDICATORS

The nature of a substance (acidic or alkaline) is identified by using the acid-base indicator.

Acid - base indicators are a class of compounds that show distinctly different colours in acidic and alkaline solutions.

Indicator	Acidic	Alkaline	pH range
Litmus	Red	Blue	5.0 – 8.0
Methyl orange	Red	Yellow	3.1 - 4.4
Phenolphthalein	Colorless	Pink	8.3 – 10.0
Turmeric	Yellow	Orange	8.0 – 10.0

### Litmus

Litmus is a dye which is obtained from lichens, and it is used as an indicator for determining the acidic and alkaline nature of the given solution. It shows red colour in acidic solutions and blue colour in alkaline solutions.

*When we eat sugar rich foods, acid is formed inside our mouth as the bacteria present there break the sugar in Lactic acid. This decreases the pH of our mouth resulting in tooth decay particularly when the pH goes below 5.8*

A strip of filter paper soaked in litmus solution and dried, called litmus paper, is more commonly used. If a blue litmus paper turns red in a solution, the solution is acidic, and if a red litmus paper turns blue in a solution, then it is an alkaline solution.

### Methyl Orange

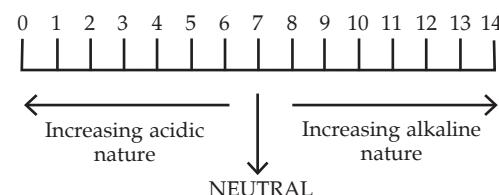
In an acid - base titration, methyl orange indicator is used as an indicator to determine the end point of the titration. When the indicator is added to the base the alkali turns into yellow, by adding the acid drop by drop we can see the colour change from yellow to orange at one particular point. This point is called neutralization point or end point for the given acid -

base reaction. If we add one more drop of acid to the alkali, the solution becomes red indicating that excess of acid is added to the alkali solution.

In higher classes we will study the acid-base concentration using titration method.

## 3.19. UNIVERSAL INDICATORS AND pH PAPER

To find the pH of the solution the pH paper is used and it consists of different colours at different pH values. pH papers use universal indicators for acidic and alkaline solutions.



The normal pH scale ranges from 0 to 14 as shown above.

### The pH Scale

Danish biochemist Sorenson proposed the pH scale to measure the acidic or basic character of a solution. The term pH was derived from the French word "puissance de hydrogen or pouvoir hydrogen", – concentration of hydrogen.

The pH scale is a very convenient method of expressing acidity and alkalinity of a solution and is defined as follows.

*"The pH of a solution is the logarithm of the inverse of the molar concentration of hydrogen ions ( $\text{H}^+$ ) in the solution. It can also be expressed as the negative logarithm of the molar concentration of hydrogen ions in the solution".*

Mathematical representation of pH;

$$\text{pH} = \log_{10} \frac{1}{[\text{H}^+]} = -\log_{10} [\text{H}^+]$$

Pure water is neutral and ionizes very slightly to yield equal number of  $\text{H}^+$  ( $\text{H}_3\text{O}^+$ ) and  $\text{OH}^-$  ions and it is represented as  $K_w$ .

For pure water at 25°C,

The ionic product of water ( $K_w$ ) =  $[\text{H}^+] [\text{OH}^-] = 1 \times 10^{-14}$

As the concentrations of  $\text{H}^+$  and  $\text{OH}^-$  are equal in pure water,

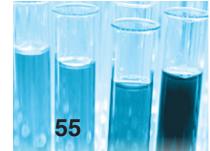
$$[\text{H}^+] = [\text{OH}^-] = 10^{-7}$$

Hence the pH of pure water is given by

$$\text{pH} = \log_{10} = \log_{10} 10 \cdot 10^{-7} = 7.$$

Hence,

- ◆ The pH of pure water is 7.



- ◆ The pH scale ranges from 0 – 14.
- ◆ The pH of an acidic solution is less than 7.
- ◆ The pH of an alkaline solution is more than 7.
- ◆ The pH of a neutral solution is 7.

Substance	pH Value
Gastric juice	1.0 – 3.0
Soft drinks	2.0 – 4.0

Lemon juice	2.2 – 2.4
Vinegar	2.4 – 3.4
Pure water	7.0
Blood	7.3 – 7.5
Egg white	7.6 – 8.0
Milk of magnesia	7.6 – 8.0
Sodium hydroxide(0.1N)	13.0

Colour Changes by Universal Indicators															
pH range	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
Colour Change		RED		ORANGE			YELLOW GREEN		BLUE			INDIGO		VIOLET	

### Universal Indicator and pH

#### 3.20 ACID RAIN

Natural rainwater is not pure water as it contains dissolved  $\text{CO}_2$ ,  $\text{SO}_2$ , etc. Hence its pH is not 7 but about 5.6, i.e., slightly acidic.



A low concentration of  $\text{H}_3\text{O}^+$  gives rainwater a pH less than 7. But, in case of acid rain the pH is much lower as the rainwater contains dissolved  $\text{H}_2\text{SO}_4$  or  $\text{HNO}_3$  from the burning fossil fuels (petroleum, coals etc.) containing sulphur. The fossil fuels are burned in power plants to generate electricity. Large amount of  $\text{SO}_2$  is formed and released into the atmosphere.

When these sulphur oxide fumes come in contact with rain water, they form  $\text{H}_2\text{SO}_4$ , which acidifies the rainwater. The rain water, thus formed is called acid rain, having a pH lower than 5.6. Acid rain destroys forest vegetation and makes lakes and streams too acidic for fish and shellfish to survive. Overall it destroys the natural environment.

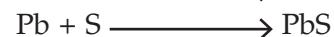
Acid rain also affects statues and buildings built of marble ( $\text{CaCO}_3$ ) as the  $\text{H}_2\text{SO}_4$  present in it dissolves the  $\text{CaCO}_3$  of the marble.

#### 3.21. GENERAL METHODS FOR PREPARING SALTS

- ◆ Direct Combination
- ◆ Decomposition
- ◆ Double Decomposition
- ◆ Displacement
- ◆ Neutralization of Acid and Base

#### Direct Combination

In this method a metal is heated with a non-metal or a halogen is passed over it to form a salt.



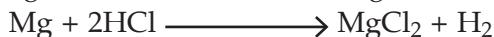
#### Decomposition

Salts of carbonates, hydrogen carbonates and basic carbonates of metals react with dilute acids to give the corresponding salt, and  $\text{CO}_2$  is formed with effervescence.



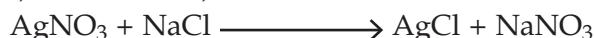
#### Displacement

In this method, salts are prepared by the reaction of dilute mineral acids with active metals.



#### Double Decomposition

When the solutions of two metal salts are mixed, the anion of one combines with the cation of the other, or vice versa, to form two new salts.



**Neutralization of acid and base :** When acid and base get neutralized the result is the formation of salt and water.



### Water of Crystallization

The amount of water bound loosely to the salt molecule during its crystallization in aqueous solution. Salts containing water of crystallization are known as hydrated salts. For example,  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ ,  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ .

### LABORATORY PREPARATION OF SOME IMPORTANT SALTS

#### Preparation of Sodium Sulphate ( $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ )

##### Principle :

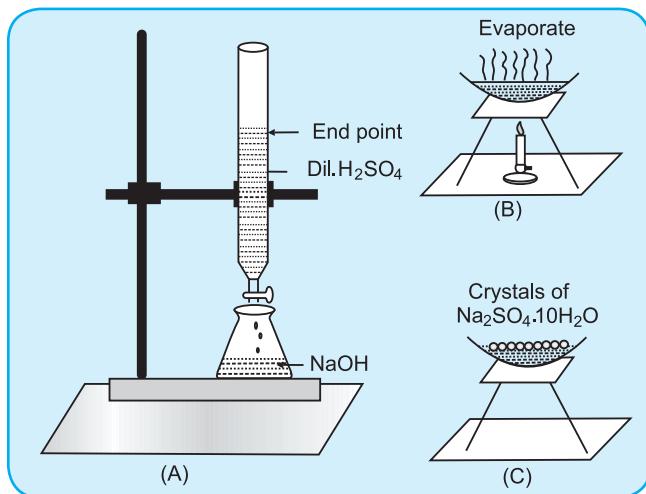
Sodium sulphate can be obtained in the laboratory by the action of sulphuric acid on sodium hydroxide through volumetric analysis.

##### Reaction :



##### Procedure :

A known amount of sodium hydroxide is taken and sulphuric acid is added to neutralize it. The neutralization point is then noted :



- ◆ Take sulphuric acid in the burette up to the zero mark and take 25 ml pipette and pipette out sodium hydroxide. Pour it in the clean conical flask.
- ◆ Add 2-3 drops of indicator (phenolphthalein) into the conical flask containing sodium hydroxide. Now the solution turns into pink colour.
- ◆ Titrate the sulphuric acid present in the burette against the sodium hydroxide present in the conical flask.
- ◆ Titrate drop by drop till the pink colour solution turns to colourless and stop the titration immediately.
- ◆ The point at which the colour changes from pink to colourless is called the neutralization point or end point.

- ◆ Transfer the conical flask solution into a china dish and heat it till the point of crystallization.
- ◆ Allow the hot solution to cool down; now we see crystals of  $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$  appears.
- ◆ Separate the crystals from the mother liquor and dry it off using blotting paper.

#### Preparation of an Acid Salt—Sodium Bicarbonate ( $\text{NaHCO}_3$ )

##### Principle :

Sodium bicarbonate is obtained by passing carbon dioxide gas over a cold solution of sodium carbonate.

##### Reaction :



##### Procedure :

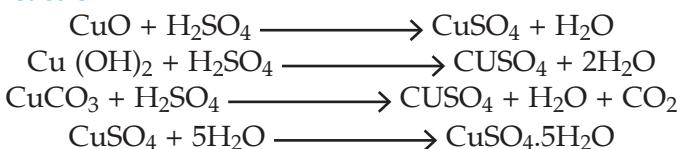
- ◆ Take 10 gms of anhydrous sodium carbonate salt and 250ml of distilled water in a beaker.
- ◆ Dissolve sodium carbonate salt in water.
- ◆ Keep this mixture under freezing condition.
- ◆ Pass the carbon dioxide gas through the glass tube into the solution.
- ◆ Crystals of sodium bicarbonate will precipitate out after some time.
- ◆ Filter the crystals using a filter paper and a funnel.
- ◆ Dry the sodium bicarbonate crystals in the filter paper.

#### Preparation of Copper (II) Sulphate (Blue Vitriol) ( $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ )

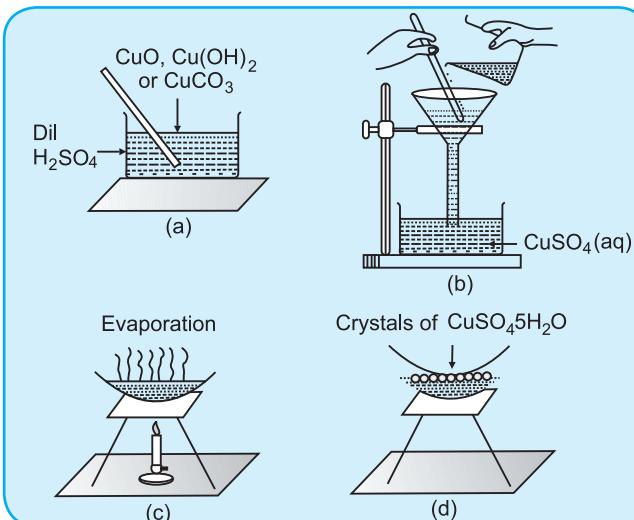
##### Principle :

Blue vitriol is obtained by the action of sulphuric acid on the oxide, hydroxide or carbonate of copper.

##### Reaction :



##### Procedure :





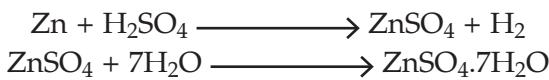
- ◆ Take 50 ml of sulphuric acid in a 250 ml beaker and heat the acid.
- ◆ Add some CuO into the hot sulphuric acid by stirring the solution.
- ◆ The salt, which is added is dissolved slowly by constant stirring with glass rod.
- ◆ When no more salt is dissolved, stop adding salt.
- ◆ Filter the hot solution containing the residue using a filter paper in a funnel and collect the filtrate in a China dish.
- ◆ Take the filtrate and heat it to evaporate the water till the point of crystallization.
- ◆ Allow the solution to cool and now we can find the formation of blue colour crystals of copper sulphate.
- ◆ Dry the crystals using the blotting paper.

### Preparation of Zinc Sulphate (White Vitriol ( $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$ ))

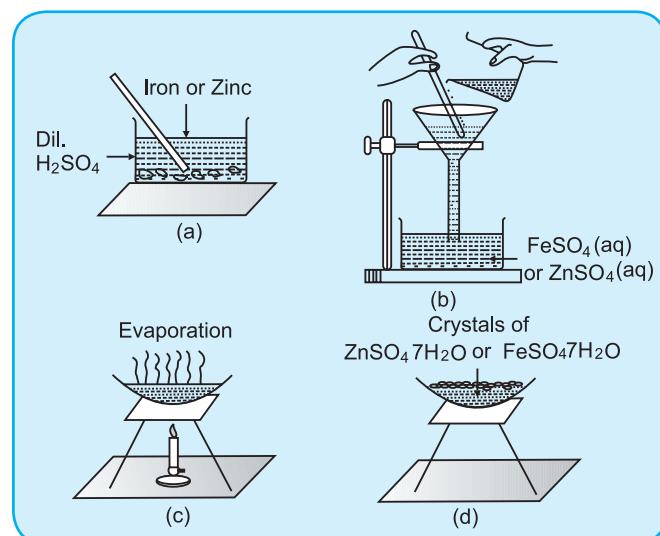
#### Principle :

White vitriol is obtained in the laboratory by the displacement of hydrogen in an acid by the metal zinc.

#### Reaction :



#### Procedure :



- ◆ Take 50 ml of dilute sulphuric acid in a 350 ml beaker and heat it.
- ◆ Add some zinc granules and stir constantly, using a glass rod.
- ◆ Add zinc till the effervescence (bubbling of hydrogen gas) ceases.

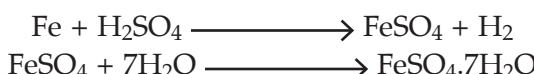
- ◆ Filter off the unused zinc using a funnel.
- ◆ Collect the filtrate in a China dish and heat.
- ◆ Heat the solution and allow the water to evaporate till the point of crystallization.
- ◆ Dry the needle like white crystals formed by using the blotting paper.

### Preparation of Iron (II) Sulphate (Green Vitriol – $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$ )

#### Principle :

Green vitriol is prepared in the laboratory by the action of sulphuric acid on iron.

#### Reaction :



#### Procedure :

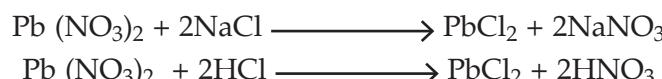
- ◆ Take 50 ml of dilute sulphuric acid in a beaker and heat it.
- ◆ Add some iron filings little by little while constantly stirring with a glass rod.
- ◆ Add till the effervescence ceases, i.e., liberation of hydrogen gas.
- ◆ Filter the solution using a funnel to remove the residues.
- ◆ Collect the filtrate in the China dish and heat the solution.
- ◆ Allow the water to evaporate till the point of crystallization.
- ◆ Dry the needle-like pale green crystals using the blotting paper.

### Preparation of Lead (II) Chloride ( $\text{PbCl}_2$ )

#### Principle :

Lead chloride can be prepared in the laboratory by double decomposition of two salts using dilute HCl or NaCl solution to obtain a solution of sodium nitrate or nitric acid. A white precipitate of lead chloride will appear.

#### Reaction :



#### Procedure :

- ◆ Take lead nitrate solution in a beaker along with an equal volume of sodium chloride solution.
- ◆ Stir the solution till a white precipitate is formed.
- ◆ Filter the precipitate to remove the sodium nitrate as filtrate.
- ◆ Transfer the residue to a China dish.

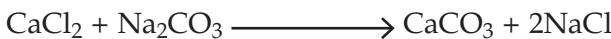
- ◆ Heat the residue in the China dish with a little amount of water.
- ◆ Upon heating, the precipitate of lead chloride will dissolve.
- ◆ Cool the solution. Needle-like crystals are formed.
- ◆ Dry the white-coloured crystals using blotting paper.

#### Preparation of Calcium Carbonate ( $\text{CaCO}_3$ ) :

##### Principle :

Calcium carbonate can be prepared by the precipitation or double decomposition reaction of sodium carbonate with calcium chloride. Calcium carbonate is obtained.

##### Reaction :



##### Procedure :

- ◆ Make calcium chloride solution in a beaker using hot water.
- ◆ Make an aqueous solution of sodium carbonate in another beaker.
- ◆ Now add the sodium carbonate solution to the hot solution of calcium carbonate.
- ◆ Add the solution little by little while constantly stirring with a glass rod.
- ◆ Add till the formation of white precipitate.
- ◆ Filter the solution and collect the white residue in a funnel.
- ◆ Wash the residue with cold water.
- ◆ Collect it in a china dish and allow to dry.
- ◆ Pure amorphous powder of calcium carbonate is obtained.

#### Preparation of Iron (III) Chloride ( $\text{FeCl}_3$ ) :

##### Principle :

Iron chloride is prepared in the laboratory by synthesis or by direct combination of elements such as chlorine gas and iron filings.

##### Reaction :



##### Procedure :

- ◆ Fill some iron filings in the combustion tube.
- ◆ Heat the combustion tube strongly.
- ◆ Through the combustion tube pass dry chlorine gas.
- ◆ The iron becomes red hot as the reaction is exothermic.
- ◆ Chlorine reacts with iron to form iron (III) chloride.

- ◆ Stop heating the combustion tube.
- ◆ Allow the vapour to dry in a cool receiver.
- ◆ The vapours are condensed as brown scales.
- ◆ Ferric chloride is highly deliquescent. Keep it dry with the help of a fused  $\text{CaCl}_2$ .

#### 3.22. DECOMPOSITION OF BICARBONATES, CARBONATES, CHLORIDES AND NITRATES BY ACIDS

The decomposition of bicarbonates, carbonates, chlorides and nitrates can be done by the following reactions.

##### Carbonate and Bicarbonate

##### Reaction

A small amount of  $\text{Na}_2\text{CO}_3$  or  $\text{NaHCO}_3$  is taken in a test tube and a few drops of dilute sulphuric acid are added.

##### Observation and Inference

A colourless gas is evolved with effervescence which is nothing but carbon dioxide gas.

The moist blue litmus turns to red colour.

The lime water turns milky on passing the evolved gas ( $\text{CO}_2$ ).

##### Chloride Salt

##### Reaction

Take small amount of  $\text{NaCl}$  in a small test tube and add a few drops of concentrated sulphuric acid and warm the solution.

##### Observation and Inference

A colourless gas with pungent smell is evolved which is the chlorine gas.

It produces thick white fumes when a glass rod moistened with concentrated ammonia is brought near it.

##### Nitrate Salt

##### Reaction

Take small amount of  $\text{KNO}_3$  in a dry test tube and add 2-3 ml of concentrated  $\text{H}_2\text{SO}_4$  to it and heat the mixture slowly.

##### Observation and Inference

A colourless or faintly brown gas is evolved which is nothing but nitric acid gas that decomposes to give  $\text{NO}_2$ .

#### 3.23. WATER OF CRYSTALLIZATION

Some salts, while crystallizing out from their solutions, unite with a definite quantity of water, which is known as the water of crystallization.

Certain salts are crystalline due to the presence of a definite number of water molecules loosely attached to the molecules of the salts. On heating, such salts lose the attached water molecules. Such water molecules are called water of crystallization or water of hydration. By heating, the salts can lose the water of crystallization partially or completely, leaving behind the anhydrous salt contained in it.

Salt	Molecular formula
Copper sulphate (Blue vitriol)	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
Ferrous sulphate (Green vitriol)	$\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$
Zinc sulphate (White vitriol)	$\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$
Magnesium sulphate (Epsom salt)	$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$
Nickel sulphate	$\text{NiSO}_4 \cdot 7\text{H}_2\text{O}$
Sodium carbonate (Washing soda)	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
Potash alum	$\text{K}_2\text{SO}_4 \cdot \text{Al}(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$

Not all crystalline salts contain water of crystallization. Examples of crystalline compounds that do not contain water of crystallization are sodium chloride and potassium chloride.

#### Crystalline Solids not containing water of Crystallization

Crystals	Formula
Common salt	$\text{NaCl}$
Nitre	$\text{KNO}_3$
Sugar	$\text{C}_{12}\text{H}_{22}\text{O}_{11}$
Potassium permanganate	$\text{KMnO}_4$

#### Experiment to show that copper sulphate crystals contain water of crystallization

Take blue copper sulphate crystals in a China dish. Heat the china dish over a flame for some time. You will be able to see water escaping from the crystals in the form of water vapour. When blue copper sulphate crystals are heated, they lose their water of crystallization and turn into a white powder. And the water of crystallization escapes as water vapour.

### 3.24. EFFLORESCENCE

Efflorescence is the property of some substances to lose wholly, or partly, their water of crystallization when their crystals are exposed to dry air even for a short time.

They become powdery. Such substances are called efflorescent substances. They lose their water of crystallization and thus their crystalline shape and become powdery when exposed to air.

Examples of efflorescent substances are :

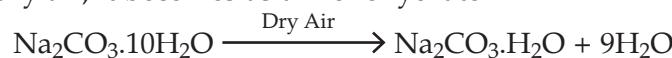
#### Glouber's Salt

Hydrated sodium sulphate  $[\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}]$  is the good example for efflorescent substance because when it is exposed to dry air, it becomes powdery anhydrous sodium sulphate.



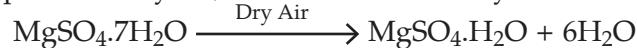
#### Washing Soda

When hydrated sodium carbonate is exposed to dry air, it becomes as a monohydrate.



#### Epsom Salt

When magnesium sulphate heptahydrate is exposed to dry air, it becomes a monohydrate.



### 3.25. HYGROSCOPIC SUBSTANCES

Certain substances absorb moisture (water vapour) from the atmosphere without dissolving in it. Such substances are called hygroscopic substances and the phenomenon is called hygroscopy.

Some examples of hygroscopic substances are :

- ◆ Concentrated sulphuric acid ( $\text{H}_2\text{SO}_4$ )
- ◆ Phosphorus pent oxide ( $\text{P}_2\text{O}_5$ )
- ◆ Quicklime ( $\text{CaO}$ )
- ◆ Silica gel

In the laboratory, these substances are used generally for drying gases, i.e., the removal of moisture from gases.

### 3.26. DELIQUESCENCE

The absorption of atmospheric water vapour by a crystalline solid until the crystal eventually dissolves into a saturated solution.

This phenomenon is called deliquescence.

#### Deliquestent substances

Examples :

Calcium chloride ( $\text{CaCl}_2$ )

Zinc chloride ( $\text{ZnCl}_2$ )

Thermodynamically deliquescence occurs when the partial pressure of the water vapour in the air

exceeds the vapour pressure of water in the saturated solution of the salt.

Then the reaction,



It occurs spontaneously.

In general, substances which are highly soluble in water have a greater tendency to be deliquescent, as concentrated solutions have a lower vapour pressure.

### NOTE

*Table salt [sodium chloride] turns moist and ultimately forms a solution on exposure to air [especially during the rainy season]. Though pure sodium chloride is not deliquescent, the commercial version of the salt contains impurities, like magnesium chloride, which is a deliquescent substance.*

This kind of impurity can be removed by passing a current of dry hydrogen chloride gas through a saturated solution of the affected salt. Pure sodium chloride is produced as a precipitate, which can be recovered by filtering and washing first with a little water and finally with alcohol.

### 3.27 DRYING AND DEHYDRATING AGENTS

#### Drying or Desiccating Agents

Certain substances remove moisture from other substances and are, therefore, called desiccants or desiccating agents or drying agents. Although the removal of large quantities of water is done by evaporation, the last traces of moisture are often held

very tightly and do not evaporate readily. Hence, they were reduced by using certain desiccating agents; almost all hygroscopic substances are desiccating agents.

#### Examples :

1. Concentrated sulphuric acid
2. Phosphorus pentoxide
3. Silica gel and quick lime

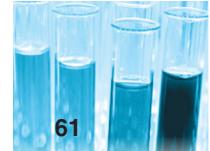
#### Dehydrating Agents

They are substances that can remove water molecules even from compounds. Concentrated sulphuric acid can remove water molecules from blue vitriol ( $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ ), so it is a dehydrating agent as well.

Drying Agent	Dehydrating Agent
They remove moisture from other substances.	They remove chemically combined elements of water in the ratio of 2 : 1 (hydrogen : oxygen) from a compound.
They are used to dry gases like chlorine, sulphur dioxide, hydrogen chloride, etc.; they are also used in desiccators to keep substances dry.	They are used to prepare substances like carbon monoxide, sugar charcoal, etc.
They represent physical change. Examples : $\text{P}_2\text{O}_5$ , fused $\text{CaCl}_2$ , $\text{CaO}$ , $\text{Con.H}_2\text{SO}_4$ .	They represent chemical change. Examples: $\text{Con.H}_2\text{SO}_4$ .

### INTEXT QUESTIONS

1. What do you understand by the term-acid ?
2. What do you understand by the term-base ?
3. What is meant by a hydronium ion ?
4. Explain the term basicity of an acid.
5. Explain the term acidity of a base.
6. Give the basicity of :
  - (a) Sulphuric acid
  - (b) Phosphoric acid
  - (c) Acetic acid
7. Name the acids present in grapes, apple and lemon.
8. Define an acidic salt.
9. Define a normal and a mixed salt.
10. What is a salt ?
11. How are the following salts prepared ?  
 $\text{FeCl}_3$ ,  $\text{PbCl}_2$ ,  $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$
12. Explain what basic, mixed, complex and double salts are.
13. What is the colour change when blue litmus is introduced into a solution of ferric chloride ?
14. What is meant by neutralization ?
15. What do you mean by water of crystallization ?
16. Define efflorescence with suitable examples.
17. Distinguish between drying and dehydrating agent.
18. Why is anhydrous calcium chloride used in desiccators?



19. Give a reason why blue vitriol changes to white when it is heated.
20. Name two substances which do not contain water of crystallization.



### POINTS TO REMEMBER

- ◆ An acid is a substance that reacts with water to form a hydronium ion.
- ◆ Bases are compounds which generally react with acid to form a salt.
- ◆ Neutralization of acids by bases results in salt formation.
- ◆ Arrhenius proposed that salts, acids and bases dissociate in aqueous solutions into ions.
- ◆ The basicity of an acid is equal to the number of replaceable hydrogen atoms present in one molecule of the acid.
- ◆ The acidity of a base is equal to the number of hydroxyl groups it loses on reaction with an acid to form a salt.
- ◆ There are six types of salts namely normal salt, acid salt, basic salt, mixed salt, double and complex salt.
- ◆ Solutions of salts of weak acids and strong bases are alkaline.
- ◆ Solutions of salts of strong acids and weak bases are acidic.
- ◆ pH of the solution is the negative logarithm of  $H^+$  ion concentration to the base ten.
- ◆ Indicators are used to find the end point of an acid-base titration.
- ◆ Deliquescence is the property of salt that absorbs the water vapour from the surrounding to form a wet salt.

### EXERCISES

#### A. FILL IN THE BLANKS

1. An acid is a compound which, when dissolved in water, gives ..... ions as the only ..... ions.
2. A ..... Salt is one in which the hydrogen of an acid has been partially replaced by .....
3. A solution X turns blue litmus red, so it must contain :
  - (a) ..... ions, another solution Y turns red litmus blue and therefore, must contain ..... ions.
  - (b) When solutions X and Y are mixed together, the products will be.....and ..... if a piece of magnesium were put into solution X, .....gas would be evolved.

#### B. MULTIPLE CHOICE TYPE QUESTIONS

Choose the correct Answer from the options given below :

1. Salts are formed when acid reacts with :
 

(a) Base	(b) Cation
(c) Water	(d) Anion
2. Tribasic acid is :
 

(a) $H_3PO_4$	(b) $HNO_3$
(c) $HCl$	(d) $H_2SO_4$
3. Weak acid among the following is :
 

(a) $CH_3COOH$	(b) $HCl$
(c) $H_2SO_4$	(d) $HNO_3$

#### 4. Complete the following reaction :

- |  |             |
|--|-------------|
| $Na_2CO_3 + 2 HCl \rightarrow NaCl + CO_2 +$ _____ |             |
| (a) $H_2O$   | (b) $H_2$   |
| (c) $NH_3$   | (d) $Na_2O$ |

#### 5. pH of an acidic solution is :

- |                   |                     |
|-------------------|---------------------|
| $(a)$ Less than 7 | $(b)$ more than 7   |
| $(c)$ equal to 7  | $(d)$ None of these |

#### 6. Green vitriol is :

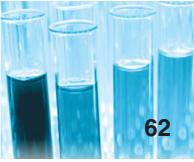
- |                          |                           |
|--------------------------|---------------------------|
| $(a) FeSO_4 \cdot 7H_2O$ | $(b) FeSO_4 \cdot 10H_2O$ |
| $(c) FeSO_4$             | $(d) FeSO_4 \cdot 3H_2O$  |

#### C. ANSWER THE FOLLOWING QUESTIONS BASED ON THE INSTRUCTIONS GIVEN BELOW

1. Arrange the following acids on the basis of increasing acidic strength.  
[ $HCl$ ,  $H_2SO_4$ ,  $HNO_3$ ]
2. Classify the following as Dibasic and Tribasic acids : [ $H_2SO_4$ ,  $H_3PO_4$ ]
3. List one use of the following acids :  
[ $HCl$ ,  $H_2SO_4$ ,  $HNO_3$ ]
4. Classify the following as Acidic and Basic salts.  
 $H_2SO_4$ ,  $Pb(OH)_2$ ,  $Cu(OH)_2$ ,  $H_2CO_3$ ,  $Mg(OH)_2$

#### D. MATCH THE COLUMNS

Column I	Column II
1. $H_2SO_4$	Monoacidic base
2. $H_3PO_4$	Less than 7



3. NaOH	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
4. Acidic solutions pH	$\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$
5. White Vitriol	Acidic salt
6. Blue Vitriol	Tribasic acid

### E. WRITE THE BALANCED CHEMICAL EQUATIONS FOR EACH OF THE FOLLOWING

- Dilute sulphuric acid and barium chloride solution.
- Dilute hydrochloric acid and lead nitrate solution.
- Magnesium sulphate solution is mixed with barium chloride solution.
- Making use only of substances chosen from those given below :

Dilute sulphuric acid, sodium carbonate  
 Zinc                                  Sodium sulphite  
 Lead                                  Calcium carbonate

Give the equations for the reactions by which you could obtain ;

- Hydrogen
  - Sulphur dioxide
  - Carbon dioxide
  - Zinc carbonate
- Write the balanced equations for the following reactions :
    - Lead sulphate from lead nitrate solution and dilute sulphuric acid.
    - Copper sulphate from copper and concentrated sulphuric acid.
    - Lead chloride from lead nitrate solution and sodium chloride solution.
    - Ammonium sulphate from ammonia and dilute sulphuric acid.
    - Sodium chloride from sodium carbonates solution and dilute hydrochloric acid.
  - The preparation of lead sulphate from lead carbonate is a two-step process. (Lead sulphate cannot be prepared by adding dilute sulphuric acid to lead carbonate).

- What is the first step that is required to prepare lead sulphate from lead carbonate ?
- Write the equation for the reaction that will take place when this first step is carried out.

### F. VERY SHORT QUESTIONS

- Mention the colour change when the following indicators are added :  
 Alkaline phenolphthalein solution and methyl orange solution, neutral litmus solution
  - Write the terms defined by the following?
    - A salt containing a metal ion surrounded by other ions or molecules
    - A base which is soluble in water
- [2008]

### G. SHORT QUESTIONS

- Define an acid.
- What is an alkali?
- Why is acetic acid a monobasic acid?
- HCl dissolves in water, forming an acidic solution. Give three distinct tests that you will carry out on this solution to illustrate the typical properties of an acid.

### H. LONG QUESTIONS

- Classify the acids based on their basicity and give complete equation for their ionisation in water.
- Write about any 3 acid-base indicators in brief.
- Explain the phenomenon of Acid rain ?
- Give the complete laboratory preparation, with the equations for Green vitrol.
- Explain the following terms with examples :
  - Efflorescence
  - Deliquescence
  - Hygroscopic agents
  - Drying agents

### RECENT BOARD QUESTIONS

#### Question 1.

- (a) Choose the correct answer from the options given below :

- Hydroxide of this metal is soluble in sodium hydroxide solution. [2011]  
 (A) Magnesium      (B) Lead  
 (C) Silver            (D) Copper
- Which of the following metallic oxides cannot be reduced by normal reducing agents ? [2012]

- Magnesium oxide (B) Copper(II) oxide
- Zinc oxide (D) Iron(III) oxide
- Which of these will act as a non-electrolyte ?  
 (A) Liquid carbon tetrachloride [2013]  
 (B) Acetic acid  
 (C) Sodium hydroxide aqueous solution acid  
 (D) Potassium chloride aqueous solution



- (iv) Which one of the following will not produce an acid when made to react with water ? [2013]  
 (A) Carbon monoxide (B) Carbon dioxide  
 (C) Nitrogen dioxide (D) Sulphur trioxide
- (v) Identify the metallic oxide which is amphoteric in nature : [2013]  
 (A) Calcium oxide (B) Barium oxide  
 (C) Zinc oxide (D) Copper(II) oxide.

**(b) Match the columns :** [2011,12,16]

List X		List Y	
(i)	$Pb^{2+}$	A.	<i>Reddish brown</i>
(ii)	$Fe^{2+}$	B.	<i>White insoluble in excess</i>
(iii)	$Zn^{2+}$	C.	<i>Dirty green</i>
(iv)	$Fe^{3+}$	D.	<i>White soluble in excess</i>
(v)	$Cu^{2+}$	E.	<i>White soluble in excess</i>
(vi)	$Ca^{2+}$	F.	<i>Blue</i>
(vii)	<i>Acid salt</i>	G.	<i>Simple displacement</i>
(viii)	<i>Double salt</i>	H.	<i>Sodium hydrogen sulphate</i>
(ix)	$ZnCO_3$ from $ZnSO_4$	I.	<i>Ferrous ammonium sulphate</i>
(x)	$Pb(NO_3)_2$ from $PbO$	J.	<i>Titration</i>
(xi)	$MgCl_2$ from $Mg$	K.	<i>Neutralization</i>
(xii)	$FeCl_3$ from $Fe$	L.	<i>Precipitation</i>
(xiii)	$NaNO_3$ from $NaOH$	M.	<i>Combination</i>

**(c) What would you observe in each of the following cases ?**

- (i) Ammonium hydroxide is first added in a small quantity and then in excess to a solution of copper sulphate. [2011]
- (ii) Crystals of washing soda when exposed to air ? Name the phenomenon exhibited. [2011,16]
- (iii) The salt ferric chloride is exposed to the atmosphere. [2016]
- (iv) Action of warm water on magnesium nitride. [2014]

**Question 2.**

- (a)** From the list given below, select the word(s) required to correctly complete blanks (i) to (v) in the following passage. The words from the list are to be used only once. Write the answers as (a) (i), (ii), (iii) and so on. Do not copy the passage. [2013]
- [ammonia, ammonium, carbonate, carbon dioxide, hydrogen, hydronium, hydroxide, precipitate, salt, water]

- (i) A solution M turns blue litmus red, so it must contain (i) \_\_\_\_\_ ions; another solution O turns red litmus blue and hence, must contain (ii) \_\_\_\_\_ ions.

- (ii) When solutions M and O are mixed together, the products will be (iii) \_\_\_\_\_ and (iv) \_\_\_\_\_ .

- (iii) If a piece of magnesium was put into a solution M, (v) \_\_\_\_\_ gas would be evolved.

**(b) Identify the gas evolved in the following reactions when :** [2013]

- (i) Sodium propionate is heated with soda lime.  
 (ii) Potassium sulphite is treated with dilute hydrochloric acid.  
 (iii) A few crystals of  $KNO_3$  are heated in a hard glass test tube.

**(c) Give suitable chemical terms for the following :** [2013]

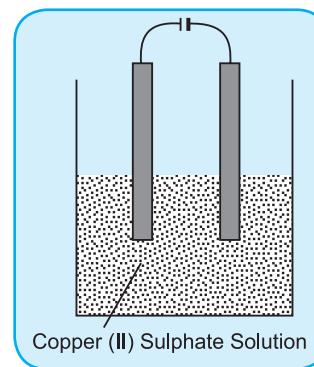
- (i) A salt formed by incomplete neutralization of an acid by a base.  
 (ii) A definite number of water molecules bound to some salts.  
 (iii) The process in which a substance absorbs moisture from the atmospheric air to become moist, and ultimately dissolves in the absorbed water.

**(d) State two relevant observations for each of the following :** [2013]

- (i) Ammonium hydroxide solution is added to copper (II) nitrate solution in small quantities and then in excess.  
 (ii) Ammonium hydroxide solution is added to zinc nitrate solution in minimum quantities and then in excess.  
 (iii) Lead nitrate crystals are heated in a hard glass test tube.

**Question 3.**

**(a) Copper sulphate solution is electrolysed using copper electrodes.**



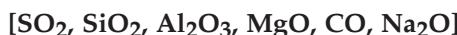
Study the diagram given below and answer the question that follows : [2013]

- (i) Which electrode to your left or right is known as the oxidising electrode and why ? [2]  
 (ii) Write the equation representing the reaction that occurs. [1]  
 (iii) State two appropriate observations for the above electrolysis reaction.

**Question 4.**

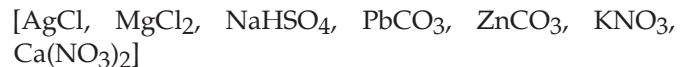
**(a) Differentiate between the terms strong electrolyte and weak electrolyte. (stating any two differences)** [2015]

(b) Choose the most appropriate answer from the following list of oxides which fit the description. Each answer may be used only once : [2015]



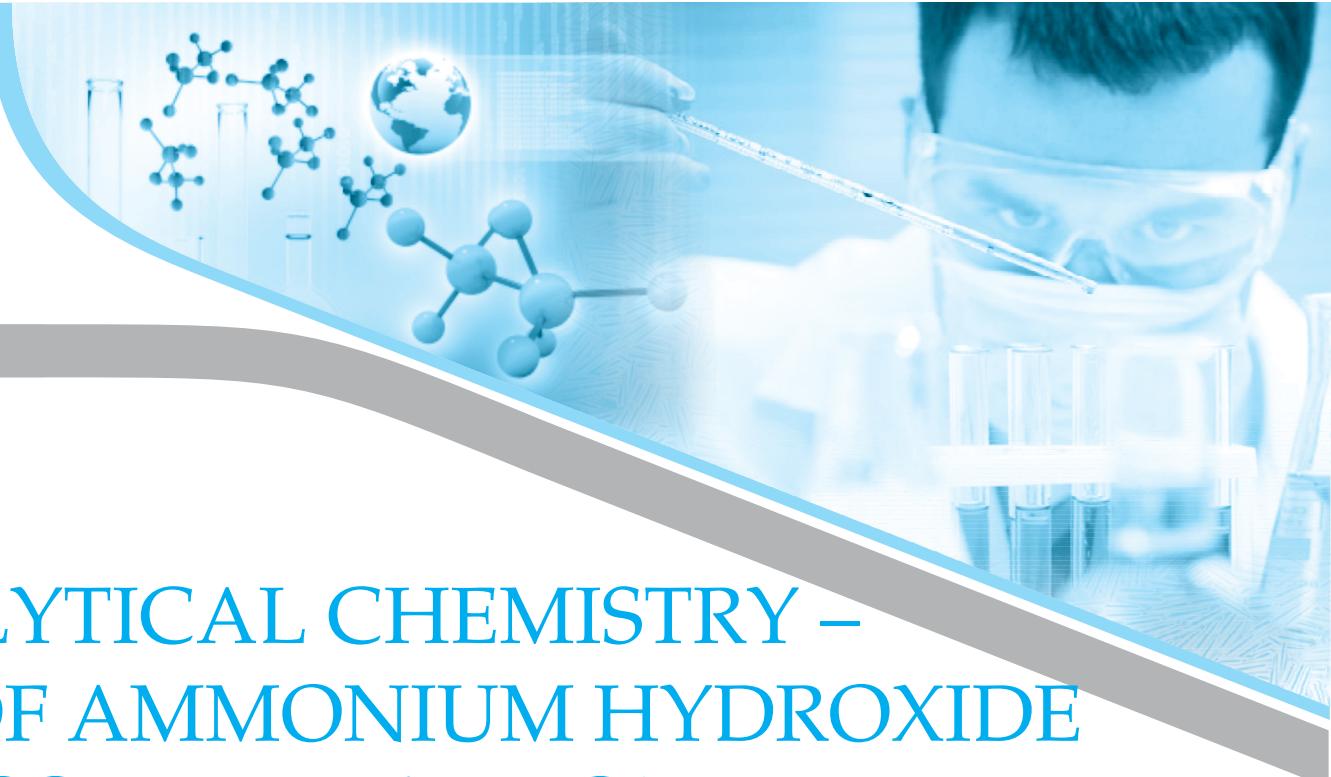
- (i) A basic oxide.
- (ii) An oxide which dissolves in water forming an acid.
- (iii) An amphoteric oxide.
- (iv) A covalent oxide of a metalloid.

(c) From the list of the following salts choose the salt that most appropriately fits the description given in the following : [2015]



- (i) A deliquescent salt.
- (ii) An insoluble chloride.
- (iii) On heating, this salt gives a yellow residue when hot and white when cold.
- (iv) On heating this salt, a brown coloured gas is evolved.





# 4

# ANALYTICAL CHEMISTRY – USE OF AMMONIUM HYDROXIDE AND SODIUM HYDROXIDE

## LEARNING OUTCOMES

- 4.0. Introduction
- 4.1. Identifying Ions by Colour

- 4.2. Use of NaOH and NH<sub>4</sub>OH in Qualitative Analysis
- 4.3. Salts and their Colours

## 4.0. INTRODUCTION

Analytical Chemistry is the branch of Chemistry that deals with identification of a given substance and assessing the quantity of different elements in the substance. For example, almost all products that we consume are subjected to analysis to affirm that they are free from contaminants. This tells us how important the field of Analytical Chemistry is. Thus, analytical chemistry can be regarded as the study of the chemical composition and structure of substances. The substances are subjected to two types of chemical analyses. They are :

- ◆ Qualitative analysis
- ◆ Quantitative analysis

'Qualitative analysis' involves *identification of an unknown substance by observing its response to certain chemical tests*. These tests are performed using various reagents, heat, etc.

*A Reagent is a substance that reacts with another substance under favourable conditions.*

For example, in qualitative analysis, two reagents NaOH, a strong alkali, and NH<sub>4</sub>OH a weak alkaline (base) solution, are widely used to determine substances such as magnesium chloride, magnesium

sulphate, calcium chloride, zinc sulphate, zinc chloride, lead sulphate, ammonium chloride, lead chloride, copper nitrate etc. Quantitative analysis involves determination of *the quantity of a substance present in a given sample*.

## 4.1. IDENTIFYING IONS BY COLOUR

In qualitative analysis, most of the observations are done visually, by the colour of the substances.

The salts of representative elements of group IA to group VII A are generally colourless, while those of the transition elements of group IB to VIIIB and group VIII are generally coloured. Some of the examples of coloured and colourless salts are given below :

**Table 1.1. Coloured Cations**

Cations	Symbol	Colour
Cupric ion	Cu <sup>2+</sup>	Blue
Ferrous ion	Fe <sup>2+</sup>	Light green
Ferric ion	Fe <sup>3+</sup>	Yellow
Nickel ion	Ni <sup>2+</sup>	Green
Chromium ion	Cr <sup>3+</sup>	Green
Manganese ion	Mn <sup>2+</sup>	Pink

**Table 1.2. Colourless cations**

Cations	Symbol
Ammonium ion	$\text{NH}_4^+$
Sodium ion	$\text{Na}^+$
Calcium ion	$\text{Ca}^{2+}$
Magnesium ion	$\text{Mg}^{2+}$
Aluminum ion	$\text{Al}^{3+}$
Lead ion	$\text{Pb}^{2+}$
Potassium ion	$\text{K}^+$

**Table 1.3. Coloured Anions**

Anion	Symbol	Colour
Permanganate ion	$\text{MnO}_4^-$	Pink
Dichromate ion	$\text{Cr}_2\text{O}_7^{2-}$	Orange
Chromate ion	$\text{CrO}_4^{2-}$	Yellow

**Table 1.4 Colourless Anions**

Anions	Symbol
Chloride ion	$\text{Cl}^-$
Carbonate ion	$\text{CO}_3^{2-}$
Sulphide ion	$\text{S}^{2-}$
Sulphate ion	$\text{SO}_4^{2-}$

Action of alkali on metal cations results in the formation of their hydroxides, which often appears as precipitates.

*Precipitation is the process of formation of a solid (precipitate) when solutions are mixed.*

Only compounds which are insoluble in water form precipitates.

#### 4.2. USE OF NAOH AND NH<sub>4</sub>OH IN QUALITATIVE ANALYSIS

The two common alkalis NaOH and NH<sub>4</sub>OH are used in identifying inorganic compounds. Since sodium hydroxide and potassium hydroxide are strong bases, they

- ◆ Give more OH ions to precipitate the insoluble metal hydroxide or oxides.
- ◆ Dissolve amphoteric metals, and metal oxides and hydroxides to form soluble salts. Since ammonium hydroxide is a weak base, it gives less OH ions to precipitate insoluble metal hydroxides or oxides.
- ◆ Interacts with insoluble metal hydroxides like Zn(OH)<sub>2</sub> and Cu(OH)<sub>2</sub> to form soluble complex compounds like [Zn(NH<sub>3</sub>)<sub>4</sub>](OH)<sub>2</sub> and [Cu(NH<sub>3</sub>)<sub>4</sub>](OH)<sub>2</sub>.

#### Acidity of Bases

Acidity of a base is defined as *the number of hydroxyl ions that one molecule of an alkali (soluble base) gives on dissociation in water. In other words it is the number of ionizable hydrogen ions (OH<sup>-</sup>) present in one molecule of a base.*

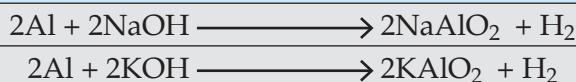
Thus, the acidity of sodium hydroxide, ammonium hydroxide and potassium hydroxide is one and they are known as mono-acid bases. The acidity of calcium hydroxide is two as a single molecule of calcium hydroxide gives two hydroxyl ions on dissociation in water. So it is called a diacid base.

#### Action of Caustic Alkali on Some Metals

Metals like aluminium, zinc, lead and stannum react with caustic alkalis like NaOH and KOH to form salts while liberating hydrogen. When these alkalis are in hot and concentrated conditions the reaction rate is rapid.

#### Table 1.5 Reactions of NaOH and KOH with Al

Aluminum reacts with NaOH and KOH to form sodium aluminate and potassium aluminate with the liberation of hydrogen gas



#### Table 1.6. Reactions of NaOH and KOH with Zn

Zinc reacts with NaOH and KOH to form sodium zincate and potassium zincate with the liberation of hydrogen gas



#### Table 1.7. Reactions of NaOH and KOH with Pb

Lead reacts with NaOH and KOH to form sodium plumbate and potassium plumbate with the liberation of hydrogen gas



Metals like magnesium, iron and copper do not react with hydroxides of sodium, potassium and ammonium.

#### Action of Caustic Alkali on Amphoteric Oxides and Hydroxides

#### Amphoteric Oxides

*Amphoteric Oxides and hydroxides are those compounds which react with both acids and alkalis to form salt and water.* In other words,

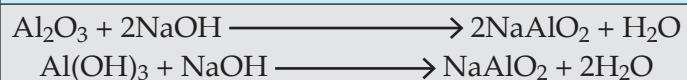
An amphoteric oxide or hydroxide of a metal is one that behaves as a base in the presence of an acid and as an acid in the presence of a strong base.

Examples :  $\text{Al}_2\text{O}_3$ ,  $\text{Al}(\text{OH})_3$ ,  $\text{ZnO}$ ,  $\text{Zn}(\text{OH})_2$ ,  $\text{PbO}$ ,  $\text{Pb}(\text{OH})_2$ ,  $\text{SnO}$  and  $\text{Sn}(\text{OH})_2$

They dissolve in aqueous  $\text{NaOH}$  as well as  $\text{KOH}$  to form soluble salts.

**Table 1.8. Reactions of Sodium Hydroxide with Aluminium and its Oxides**

**Aluminium oxide and Aluminium hydroxide react with sodium hydroxide to give the following salts and water**



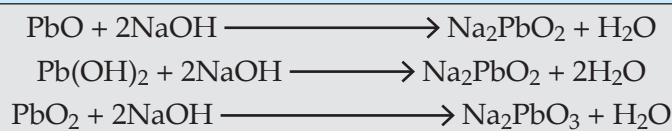
**Table 1.9 Reactions of Sodium Hydroxide with Zinc and its Oxides**

**Zinc oxide and Zinc hydroxide react with sodium hydroxide to give the following salts and water**



**Table 1.10 Reaction of sodium hydroxide with lead oxides**

**Lead oxide, Lead hydroxide and Lead dioxide react with sodium hydroxide to give the following salts and water**

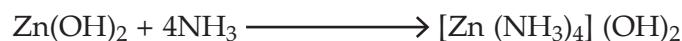


If potassium hydroxide is used in place of sodium hydroxide, then potassium salts are formed.

This property of amphoteric oxides and hydroxides distinguishes them from those that are not amphoteric, e.g.,  $\text{MgO}$ ,  $\text{Mg}(\text{OH})_2$ ,  $\text{CaO}$ ,  $\text{Ca}(\text{OH})_2$ ,  $\text{CuO}$ ,  $\text{Cu}(\text{OH})_2$ ,  $\text{Fe}_2\text{O}_3$  and  $\text{Fe}(\text{OH})_3$ .

#### Action of Ammonium Hydroxide on Metal Hydroxides

Many transition metal ions react with ammonia to form complex compounds. The hydroxides of such metals interact with  $\text{NH}_3$  solutions ( $\text{NH}_4\text{OH}$ ), forming soluble complexes.



An ammonia solution does not react with metals.

#### 4.3. SALTS AND THEIR COLOURS

**Table 1.11 Salts and their colours**

Salt	Formula	Colour
Sodium	Na	White
Calcium	Ca	White
Magnesium	Mg	White
Zinc	Zn	White
Lead	Pb	White
Lead(II) oxide	PbO	Yellow
Lead(IV) oxide	PbO <sub>2</sub>	Chocolate brown
Red lead	Pb <sub>3</sub> O <sub>4</sub>	Red or orange red
Lead sulphide	PbS	Black
Lead iodide	PbI <sub>2</sub>	Yellow
Iron(II)oxide	FeO	Black
Ferrous sulphate	FeSO <sub>4</sub> .7H <sub>2</sub> O	Green
Ferrous sulphate	FeSO <sub>4</sub>	White pale
Copper(I)oxide	Cu <sub>2</sub> O	Blue
Copper carbonate	CuCO <sub>3</sub>	Green
Copper sulphate	CuSO <sub>4</sub>	Blue
Copper sulphide	CuS	Black
Copper nitrate	Cu(NO <sub>3</sub> ) <sub>2</sub>	Green
Mercury	Hg	White
Mercury sulphide	HgS	Black
Mercury oxide	HgO	Orange red
Silver oxide	Ag <sub>2</sub> O	Grey
Silver sulphide	Ag <sub>2</sub> S	Black
Silver bromide	AgBr	Yellow

#### The action of sodium hydroxide and ammonium hydroxide on some cations

Let us discuss the action of aqueous  $\text{NaOH}$  and  $\text{NH}_4\text{OH}$  on some cations like  $\text{Mg}^{2+}$ ,  $\text{Ca}^{2+}$ ,  $\text{Fe}^{2+}$ ,  $\text{Fe}^{3+}$ ,  $\text{Cu}^{2+}$ ,  $\text{Zn}^{2+}$ ,  $\text{Pb}^{2+}$  and  $\text{NH}_4^{+}$ .

To study this reaction, 2 grams of the sample salt is dissolved in 3-5 ml of water to make a salt solution.

### Action on $Mg^{2+}$

**Table 1.12. Reactions of Magnesium with Sodium Hydroxide and Ammonium Hydroxide**

Experiment	Reaction	Observation
An aqueous solution of <b>sodium hydroxide</b> is added to an aqueous solution of Epsom salt.	$MgSO_4 + 2NaOH \longrightarrow Mg(OH)_2 + Na_2SO_4$	A white precipitate of $Mg(OH)_2$ is obtained that remains constant with the addition of excess reagent.
An aqueous solution of <b>ammonium hydroxide</b> is added to an aqueous solution of Epsom salt.	$MgSO_4 + 2NH_4OH \longrightarrow Mg(OH)_2 + (NH_4)_2SO_4$	A white precipitate of $Mg(OH)_2$ is formed, which is unaffected by excess of the reagent.

### Action on $Ca^{2+}$

**Table 1.13. Reactions of Calcium with Sodium Hydroxide and Ammonium Hydroxide**

Experiment	Reaction	Observation
An aqueous solution of <b>sodium hydroxide</b> is added to the concentrated solution of calcium chloride.	$CaCl_2 + 2NaOH \longrightarrow Ca(OH)_2 + 2NaCl$	A white precipitate of $Ca(OH)_2$ is obtained, which remains insoluble in excess of the reagent.
An aqueous solution of <b>ammonium hydroxide</b> is added to the concentrated solution of calcium chloride.	$CaCl_2 + 2NH_4OH \longrightarrow Ca(OH)_2 + 2NH_4Cl$	A white precipitate of $Ca(OH)_2$ is formed, which is unaffected by excess of the reagent.

### Action of $Zn^{2+}$ (White Vitriol)

**Table 1.14. Reactions of Zinc Sulphate with Sodium Hydroxide and Ammonium Hydroxide**

Experiment	Reaction	Observation
An aqueous solution of <b>sodium hydroxide</b> is added to an aqueous solution of white vitriol (zinc sulphate).	$ZnSO_4 + 2NaOH \longrightarrow Zn(OH)_2 + Na_2SO_4$ $Zn(OH)_2 + 2NaOH \longrightarrow Na_2ZnO_2 + 2H_2O$	A white precipitate of $Zn(OH)_2$ is obtained. Being amphoteric, $Zn(OH)_2$ dissolves in an excess of the $NaOH$ solution to form sodium zincate and water.
An aqueous solution of <b>ammonium hydroxide</b> is added to an aqueous solution of white vitriol (zinc sulphate).	$ZnSO_4 + 2NH_4OH \longrightarrow Zn(OH)_2 + (NH_4)_2SO_4$ $Zn(OH)_2 + 4NH_3 \longrightarrow [Zn(NH_3)_4](OH)_2$	A white precipitate of $Zn(OH)_2$ is formed that dissolves in excess of the $NH_3$ solution to form a soluble colourless ammine complex (Tetraamminenitrate (II) hydroxide).

### Action of $Pb^{2+}$

**Table 1.15. Reactions of Lead Nitrate with Sodium Hydroxide and Ammonium Hydroxide**

Experiment	Reaction	Observation
An aqueous solution of <b>sodium hydroxide</b> is added to an aqueous solution of lead nitrate.	$Pb(NO_3)_2 + 2NaOH \longrightarrow Pb(OH)_2 + 2NaNO_3$ $Pb(OH)_2 + 2NaOH \longrightarrow Na_2PbO_2 + 2H_2O$	A white precipitate of $Pb(OH)_2$ is formed. As it is amphoteric, it dissolves in excess of the reagent to form sodium plumbite and water.

An aqueous solution of <b>ammonium hydroxide</b> is added to an aqueous solution of lead nitrate.	$\text{Pb}(\text{NO}_3)_2 + 2\text{NH}_4\text{OH} \rightarrow \text{Pb}(\text{OH})_2 + 2\text{NH}_4\text{NO}_3$	A white precipitate of $\text{Pb}(\text{OH})_2$ is formed, which remains insoluble in excess of the reagent.
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**Action on  $\text{Fe}^{2+}$  (Green vitriol)****Table 1.16 Reactions of Ferrous Sulphate with Sodium Hydroxide and Ammonium Hydroxide**

Experiment	Reaction	Observation
An aqueous solution of <b>sodium hydroxide</b> is added to an aqueous solution of green vitriol (ferrous sulphate).	$\text{FeSO}_4 + 2\text{NaOH} \rightarrow \text{Fe}(\text{OH})_2 + \text{Na}_2\text{SO}_4$ $4\text{Fe}(\text{OH})_2 + \text{O}_2 + 2\text{H}_2\text{O} \rightarrow 4\text{Fe}(\text{OH})_3$	A dirty green precipitate of $\text{Fe}(\text{OH})_2$ is obtained, which is not soluble in excess of the reagent. The precipitate slowly turns brown due to the atmospheric oxidation of $\text{Fe}(\text{OH})_2$ into $\text{Fe}(\text{OH})_3$ .
An aqueous solution of <b>ammonium hydroxide</b> is added to an aqueous solution of green vitriol (ferrous sulphate).	$\text{FeSO}_4 + 2\text{NH}_4\text{OH} \rightarrow \text{Fe}(\text{OH})_2 + (\text{NH}_4)_2\text{SO}_4$ $4\text{Fe}(\text{OH})_2 + \text{O}_2 + 2\text{H}_2\text{O} \rightarrow 4\text{Fe}(\text{OH})_3$	A dirty green precipitate of $\text{Fe}(\text{OH})_2$ is formed, which is not soluble in excess of the reagent. The precipitate slowly turns brown due to the atmospheric oxidation of $\text{Fe}(\text{OH})_2$ into $\text{Fe}(\text{OH})_3$ .

**Action on  $\text{Fe}^{3+}$  (Ferric Chloride)****Table 1.17. Reactions of ferric chloride with sodium hydroxide and ammonium hydroxide**

Experiment	Reaction	Observation
An aqueous solution of <b>sodium hydroxide</b> is added to an aqueous solution of ferric chloride (dil HCl is added to make the solution clear).	$\text{FeCl}_3 + 3\text{NaOH} \rightarrow \text{Fe}(\text{OH})_3 + 3\text{NaCl}$	A reddish brown precipitate of $\text{Fe}(\text{OH})_3$ is obtained, which remains unchanged when excess of the reagent is added.
An aqueous solution of <b>ammonium hydroxide</b> is added to an aqueous solution of ferric chloride.	$\text{FeCl}_3 + 3\text{NH}_4\text{OH} \rightarrow \text{Fe}(\text{OH})_3 + 3\text{NH}_4\text{Cl}$	A dirty brown precipitate of $\text{Fe}(\text{OH})_3$ is formed, which remains unaffected by excess of the reagent.

**Action of  $\text{Cu}^{2+}$** **Table 1.18. Reactions of Copper Sulphate with Sodium Hydroxide and Ammonium Hydroxide**

Experiment	Reaction	Observation
An aqueous solution of <b>sodium hydroxide</b> is added to an aqueous solution of copper sulphate (blue vitriol).	$\text{CuSO}_4 + 2\text{NaOH} \rightarrow \text{Cu}(\text{OH})_2 + \text{Na}_2\text{SO}_4$ $\text{Cu}(\text{OH})_2 \rightarrow \text{CuO} + 2\text{H}_2\text{O}$	A light blue precipitate of $\text{Cu}(\text{OH})_2$ is obtained, which remains unaffected by excess of the reagent. On heating, the light blue precipitate turns black due to the formation of $\text{CuO}$ .

An aqueous solution of <b>ammonium hydroxide</b> is added to an aqueous solution of copper sulphate (blue vitriol).	$\text{CuSO}_4 + 2\text{NH}_4\text{OH} \rightarrow \text{Cu}(\text{OH})_2 + (\text{NH}_4)_2\text{SO}_4$ $\text{Cu}(\text{OH})_2 + 4\text{NH}_3 \rightarrow [\text{Cu}(\text{NH}_3)_4](\text{OH})_2$	A light blue precipitate of $\text{Cu}(\text{OH})_2$ is first obtained. The light blue precipitate of $\text{Cu}(\text{OH})_2$ readily dissolves in an excess of the $\text{NH}_3$ solution to form an intense blue solution of $[\text{Cu}(\text{NH}_3)_4](\text{OH})_2$ .
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### Action of $\text{NH}_4^+$

**Table 1.19. Reactions of Ammonium Salt with Sodium Hydroxide and Ammonium Hydroxide**

Experiment	Reaction	Observation
An aqueous solution of <b>sodium hydroxide</b> is added to an aqueous solution of ammonium chloride and boiled.	$\text{NH}_4\text{Cl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} + \text{NH}_3$ $\text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s)$	Ammonia gas is liberated, which can be detected by the following methods : <ul style="list-style-type: none"> <li>• It is a colourless gas with a characteristic smell</li> <li>• It turns moist red litmus paper blue</li> <li>• It produces dense white fumes when a glass rod moistened with concentrated HCl is brought near to it</li> <li>• The HCl that evaporates from the moistened rod reacts with <math>\text{NH}_3</math> gas to produce solid <math>\text{NH}_4\text{Cl}</math>, which appears as fumes</li> </ul>
An aqueous solution of ammonium hydroxide is added to an aqueous solution of ammonium chloride.	<b>No reaction</b>	

### INTEXT QUESTIONS

1. What do you mean by chemical analysis ?
2. Define quantitative analysis.
3. What is a precipitate ?
4. In which group do coloured salts belong in a periodic table ?
5. What are the colours of the following salts : Zinc, lead iodide and copper nitrate



### POINTS TO REMEMBER

- ◆ There are two types of analysis : qualitative analysis and quantitative analysis.
- ◆ Qualitative analysis involves identifying an unknown substance by observing its response to certain tests.
- ◆ Quantitative analysis involves determining the quantity of a substance in a given sample.
- ◆ Reagent is a substance that reacts with another substance.
- ◆ Precipitation is the process of formation of a solid (precipitate) when solutions are mixed.
- ◆ Acidity of a base is defined as the number of hydroxyl ions that one molecule of an alkali (soluble base) gives on dissociation in water.
- ◆ Amphoteric oxides and hydroxides are those compounds which can react with both acids and alkalis to form salt and water.



- ◆ Since sodium hydroxide and potassium hydroxide are strong bases, they give more  $\text{OH}^-$  ions to precipitate the insoluble metal hydroxide or oxides. They dissolve amphoteric metals, metal oxides and hydroxides to form soluble salts.
- ◆ Since ammonium hydroxide is a weak base, it gives less  $\text{OH}^-$  ions to precipitate insoluble metal hydroxides/oxides.
- ◆ They interact with insoluble metal hydroxides like  $\text{Zn}(\text{OH})_2$  and  $\text{Cu}(\text{OH})_2$  to form soluble complex compounds like  $[\text{Zn}(\text{NH}_3)_4](\text{OH})_2$  and  $[\text{Cu}(\text{NH}_3)_4](\text{OH})_2$ .

## EXERCISES

### A. FILL IN THE BLANKS

1. The transition metal ions react with ammonia to form a .....
2. ....is added to ferric chloride solution to obtain a clear solution.
3. When an ammonium salt is heated with  $\text{NaOH}$  solution..... is evolved.
4. A white precipitate is soluble in excess  $\text{NH}_4\text{OH}$  is

### B. MULTIPLE CHOICE TYPE QUESTIONS

**Choose the Correct Answer from the options given below :**

1. The solutions of which of the following salts should be colourless ?
 

(a) $\text{FeSO}_4$	(b) $\text{CuSO}_4$
(c) $\text{FeCl}_3$	(d) $\text{Zn}(\text{NO}_3)_2$
2. The solution of which of the following salts will become a deep inky blue when an excess of  $\text{NH}_4\text{OH}$  solution is added to it ?
 

(a) $\text{Cu}(\text{NO}_3)_2$	(b) $\text{FeSO}_4$
(c) $\text{FeCl}_3$	(d) $\text{Pb}(\text{NO}_3)_2$
(e) $\text{MgSO}_4$	(f) $\text{ZnCl}_2$ [ICSE 2005]
3. Which of the following metal oxides can react with acids as well as strong alkalis ?
 

(a) $\text{Ag}_2\text{O}$	(b) $\text{CuO}$
(c) $\text{Al}_2\text{O}_3$	(d) $\text{CaO}$

[ICSE 2009]

### C. CHOOSE THE MOST APPROPRIATE ANSWER FROM THE FOLLOWING FROM THE LIST OF OXIDES WHICH FIT THE DESCRIPTION. EACH ANSWER MAY BE USED ONLY ONCE

$[\text{SO}_2, \text{SiO}_2, \text{Al}_2\text{O}_3, \text{MgO}, \text{CO}, \text{Na}_2\text{O}]$

- (i) A basic oxide.
- (ii) An oxide which dissolves in water forming an acid.
- (iii) An amphoteric oxide.
- (iv) A covalent oxide of a metalloid

[2015]

### D. MATCH THE COLUMNS

Column I	Column II
1. Cupric ion	White
2. Nickel ion	Light yellow
3. Mercury salt	Green
4. Silver salt	Pink
5. Manganese ion	Grey
6. Ferrous ion	Blue
7. Ammonium hydroxide solution	Contains only molecules
8. Carbon tetrachloride	Contains ions and molecules

### E. ANSWER THE FOLLOWING BASED ON THE OBSERVATIONS

- (1) The following table shows the tests a student performed on four different aqueous solutions which are X, Y, Z and W. Based on the observations provided, identify the cation present :  
To solution X, ammonium hydroxide is added in minimum quantity first and then in excess. A dirty white precipitate is formed which dissolves in excess to form a clear solution.
  - (i) To solution Y ammonium hydroxide is added in minimum quantity first and then in excess. A pale blue precipitate is formed which dissolves in excess to form a clear inky blue solution.
  - (ii) To solution W a small quantity of sodium hydroxide solution is added and then in excess. A white precipitate is formed which remains insoluble.
  - (iii) To a salt Z calcium hydroxide solution is added and then heated. A pungent smelling gas turning moist red litmus paper blue is obtained. [2015]
- (2) Choose the salt that most appropriately fits the description given in the following :

[ $\text{AgCl}$ ,  $\text{MgCl}_2$ ,  $\text{NaHSO}_4$ ,  $\text{PbCO}_3$ ,  $\text{ZnCO}_3$ ,  $\text{KNO}_3$ ,  $\text{Ca}(\text{NO}_3)_2$ ]

- (i) A deliquescent salt.
- (ii) An insoluble chloride.
- (iii) On heating, this salt gives a yellow residue when hot and white when cold.
- (iv) On heating this salt, a brown coloured gas is evolved.

#### F. VERY SHORT QUESTIONS

1. What happens when a glass rod moistened with concentrated  $\text{HCl}$  is placed in the gas ?  
[ICSE 1995, 2001]
2. What will you observe when an excess of  $\text{NaOH}$  solution is added to a  $\text{ZnSO}_4$  solution ?  
[ICSE 1996]
3. A  $\text{NaOH}$  solution can be used to distinguish between iron (II) sulphate (or chloride) and iron (III) sulphate (or chloride) solutions. Give the colour of the precipitate formed with each solution.  
[ICSE 1995, 1998]
4. What will you observe when  $\text{NH}_4\text{OH}$  is added to an iron (III) sulphate solution ?  
[ICSE 2001]

#### G. SHORT QUESTIONS

1. What is meant by qualitative analysis ?
2. What is meant by a reagent ?
3. Define acidity of a base.
4. What do you understand by the term – amphoteric oxide and hydroxide ?
5. Find the odd one out and explain your choice. Note that valency is not a criterion.  
 $\text{Al(OH)}_3$ ,  $\text{Pb(OH)}_2$ ,  $\text{Mg(OH)}_2$ ,  $\text{Zn(OH)}_2$   
[ICSE 2009]

#### H. LONG QUESTIONS

1. Using  $\text{NaOH}$  solution, how would you distinguish.
  - (a)  $(\text{NH}_4)_2\text{SO}_4$  solution from  $\text{Na}_2\text{SO}_4$  solution
  - (b)  $\text{Zn}(\text{NO}_3)_2$  solution from  $\text{Ca}(\text{NO}_3)_2$  solution
  - (c)  $\text{FeCl}_2$  solution from  $\text{FeCl}_3$  solution
  - (d)  $\text{Pb(OH)}_2$  from  $\text{Mg(OH)}_2$ ? [ICSE 1993, 1999]
2. There are three test tubes—one contains a  $\text{Ca}(\text{NO}_3)_2$  solution, another  $\text{Zn}(\text{NO}_3)_2$  solution and the third a  $\text{Pb}(\text{NO}_3)_2$  solution. Each solution is divided into two parts. Describe what happens when
  - (a) Excess of  $\text{NaOH}$  solution is added to each portion of the reagent.
  - (b)  $\text{NH}_4\text{OH}$  is added to each portion till the reagent is in excess. [ICSE 1997]
3. How would you distinguish between  $\text{Zn}^{2+}$  and  $\text{Pb}^{2+}$  using  $\text{NH}_4\text{OH}$ ? [ICSE 2003]
4. Discuss the action of  $\text{NaOH}$  solution on solutions of  $\text{ZnSO}_4$  and  $\text{Pb}(\text{NO}_3)_2$ .
5. Give one test each to distinguish between the following pairs of chemicals :
  - (a) Zinc nitrate solution and calcium nitrate solution.
  - (b) Sodium nitrate solution and sodium chloride solution.
  - (c) Iron(III) chloride solution and copper chloride solution. [ICSE 2006]
6. Discuss the action of  $\text{NaOH}$  solution on solutions of  $\text{ZnSO}_4$  and  $\text{CuSO}_4$ .

#### RECENT BOARD QUESTIONS

##### Question 1.

- (a) Write the balanced chemical equation for each of the following reactions :

- (i) Lead nitrate solution is added to sodium chloride solution. [2011]
- (ii) Zinc oxide dissolves in sodium hydroxide.. [2011]
- (iii) Silver nitrate solution and Sodium chloride solution. [2012]
- (iv) Action of warm water on  $\text{AlN}$ . [2016]
- (v) Action of water on calcium carbide. [2016]

- (b) State your observations :

- (i) When Barium chloride solution is mixed with sodium sulphate solution. [2016]
- (ii) When ammonium hydroxide solution is added drop by drop and then in excess to each of the following solutions : [2016]

(a) Copper sulphate solution.

(b) Zinc sulphate solution.

- (c) When an aqueous solution of sodium chloride conducts electricity. [2012]

##### Question 2.

- (a) (i) Identify the cations in each of the following case : [2016]

- (a)  $\text{NaOH}$  solution when added to the solution (A) gives a reddish brown precipitate.
- (b)  $\text{NH}_4\text{OH}$  solution when added to the solution (B) gives white ppt which does not dissolve in excess.
- (c)  $\text{NaOH}$  solution when added to solution (C) gives white ppt which is insoluble in excess.



- (ii) The following table shows the tests a student performed on four different aqueous solutions which are X, Y, Z and W. Based on the observations provided, indentify the cation present : [2015]

Chemical test	Observation	Conclusion
To solution X, ammonium hydroxide is added in minimum quantity first and then in excess.	A dirty white precipitate is formed which dissolves in excess to form a clear solution.	(i)
To solution Y ammonium hydroxide is added in minimum quantity first and then in excess.	A pale blue precipitate is formed which dissolves in excess to form a clear inky blue solution.	(ii)
To solution W a small quantity of sodium hydroxide solution is added and then in excess.	A white precipitate is formed which remains insoluble.	(iii)
To a salt Z calcium hydro-xide solution is added and then heated.	A pungent smelling gas turning moist red litmus paper blue is obtained.	(iv)

- (b) Identify the anion present in the following compounds : [2012]

- (i) Compound X on heating with copper turnings and concentrated sulphuric acid liberates a reddish brown gas.
- (ii) When a solution of compound Y is treated with silver nitrate solution a white precipitate is obtained which is soluble in excess of ammonium hydroxide solution.
- (iii) Compound Z which on reacting with dilute sulphuric acid liberates a gas which turns lime water milky, but the gas has no effect on acidified potassium dichromate solution.
- (iv) Compound L on reacting with Barium chloride solution gives a white precipitate insoluble in dilute hydrochloric acid or dilute nitric acid.

- (c) Name the gas evolved when the following mixtures are heated : [2016]

- (a) (i) Calcium hydroxide and Ammonium Chloride.  
(ii) Sodium Nitrite and Ammonium Chloride.
- (b) Sodium chloride will conduct electricity only in fused or aqueous solution state.

### Question 3.

- (a) Give a chemical test to distinguish between the following pairs of compounds : [2013-12]

- (i) Sodium chloride solution and sodium nitrate solution.
- (ii) Hydrogen chloride gas and hydrogen sulphide gas.
- (iii) Ethene gas and ethane gas.
- (iv) Calcium nitrate solution and zinc nitrate solution.
- (v) Carbon dioxide gas and sulphur dioxide gas.
- (vi) Ferrous nitrate and Lead nitrate



# 5

# THE MOLE CONCEPT AND STOICHIOMETRY

## LEARNING OUTCOMES

- 5.0. Introduction
- 5.1. Definition of Gas
- 5.2. The Gas Laws
- 5.3. Gay-Lussac's Law of Combining Volumes
- 5.4. Avogadro's Law
- 5.5. Relation between Vapour Density and Relative Molecular Mass of a Gas
- 5.6. Relative Atomic Mass
- 5.7. The Carbon-12 Scale

- 5.8. Fractional Atomic Masses
- 5.9. Relative Molecular Mass
- 5.10. Gram Atomic Mass
- 5.11. Gram Molecular Mass
- 5.12. Mole Concept
- 5.13. Applications of Avogadro's Hypothesis
- 5.14. Percentage Composition
- 5.15. Empirical Formula of Compound
- 5.16. Molecular Formula of a Compound
- 5.17. Chemical Equations

## 5.0. INTRODUCTION

Anything that occupies space is called matter; there are three different forms of matter: solid, liquid and gas. These three forms are known as the states of matter. Water can exist in all the three states in nature as ice, water and steam. In this chapter, we will discuss the various gas laws based on different experiments.

*The mass of 1 mole of water is only 18 g having a volume of 18 ml. This is equivalent to 3 to 4 tablespoons of water.*

## 5.1. DEFINITION OF GAS

Gas is one of the three fundamental states of matter. It is the state of matter that has neither a definite volume nor a definite shape. It expands

readily to occupy the volume of an entire vessel, in which it is confined. Therefore, the volume and shape of gas is equivalent to that of the vessel in which it is confined. All gases behave similarly under similar conditions of temperature and pressure, as expressed by the gas laws.

### Properties of Gases

#### (a) Gases are mostly transparent.

Most gases are colourless as they absorb the wavelength of visible light. Only a few gases are coloured; for example, chlorine (yellow), bromine (brown).

#### (b) Gases exert pressure in all directions.

The molecules present in the gas move here and there in random motion inside the container it is in. The molecules hit all the side walls of the container equally. This results in the gas exerting equal pressure in all directions.





(c) They occupy space, have weight and low density. They can occupy the volume of any container's shape and have weight. They have low densities due to large intermolecular spaces.

(d) They have neither definite shape nor definite volume. The intermolecular force of attraction between the gas particles is negligible. So, the intermolecular space is large. Thus, the particles are free to move about in any direction in space and occupy the entire volume of the containing vessel.

(e) They are highly compressible. The intermolecular distance can be lessened by applying pressure on the gas.

(f) Gases expand on heating or on lowering of the pressure : the distance between molecules of the gas increases and the expansion of the gas take place as the intermolecular force of attraction between the gas particles is negligible.

(g) They diffuse readily. They have large intermolecular spaces and the molecules of different gases intermingle, i.e., diffuse immediately.

## 5.2. THE GAS LAWS

There are three gas laws that focus on the effects of changes in pressure (P), temperature (T) or number of moles (n) on the gas volume (V).

**Boyle's Law and Pressure :** It states that "at constant temperature, the volume of a given mass of gas is inversely proportional to the applied pressure". In other words, pressure and volume of a gas have an inverse relationship, when temperature is held constant.

Mathematically,

$$(V \propto \frac{1}{P}); P_1 V_1 = P_2 V_2 \text{ (at constant temperature)}$$

**Charles's Law and Temperature :** It states that "the volume of a given mass of a gas is directly proportional to its absolute temperature at constant pressure". Mathematically,

$$V \propto T; V_1 T_2 = V_2 T_1 \quad (\text{at constant pressure})$$

**Gas Equation :** An ideal gas or perfect gas strictly obeys Boyle's law and Charles's law.

Therefore, an ideal gas equation is a relationship connecting the pressure, volume and temperature of a gas.

On combining the two laws, we get :

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

This means that the volume of a given mass of dry gas varies directly with temperature and inversely with pressure.

To make this proportionality into an equation, a proportionality constant, k, is used. The equation becomes :

$$V = \frac{T}{P} \times k \text{ (constant); or } \frac{PV}{T} = k$$

If the pressure of a given mass of gas changes from  $P_1$  to  $P_2$ , volume from  $V_1$  to  $V_2$  and temperature from  $T_1$  to  $T_2$ , then the gas equation becomes :

$$P_1 V_1 = P_2 V_2 = -k \text{ (constant)}$$

### NOTE

Pressure of a gas in a closed system is measured using a device called a manometer. Pressure can be expressed in different units such as newton/metre<sup>2</sup>, atmosphere, pascal and torr.

$$1 \text{ atmosphere (atm)} = 760 \text{ mm of mercury.}$$

**Standard or Normal Temperature Pressure :** Since temperature and air pressure may vary from place to place, it is vital to choose standard reference conditions for comparing the volume of gases.

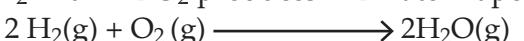
The abbreviation S.T.P or N.T.P defined by IUPAC (International Union of Pure and Applied Chemistry) denotes these specific conditions of temperature and pressure. Thus, the standard conditions for temperature and pressure are as follows :

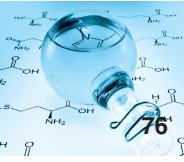
$$\text{Standard temperature} = 0^\circ\text{C} \text{ or } 273\text{K}$$

$$\begin{aligned} \text{Standard pressure} &= 760 \text{ mm of mercury} = 76 \text{ cm of} \\ &\text{Hg} = 1 \text{ atm} \end{aligned}$$

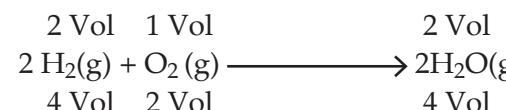
## 5.3. GAY-LUSSAC'S LAW OF COMBINING VOLUMES

In 1809, French scientist Joseph Gay-Lussac conducted an experiment in which he measured the volume of gases reacting with one another to form gaseous products. He found that at constant temperature and pressure the volume of reacting gases were always in the ratios of small whole numbers. This is known as the **law of combining volumes**. For example, the reaction of 2L H<sub>2</sub> with 1L O<sub>2</sub> produces 2 L water vapour. Similarly, the reaction of 4 L H<sub>2</sub> with 2 L O<sub>2</sub> produces 4 L water vapour.





## CERTIFICATE CHEMISTRY-X



This law is valid only for gases since the volumes of solids and liquids are considered to be zero compared to gases.

He noted similar simple relationships with other gases also. Some of the examples involving gases or vapours of other gases are given in Table 5.1 :

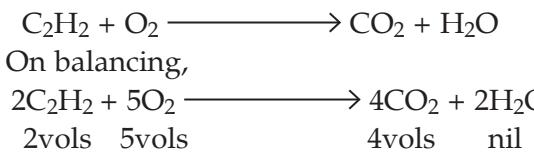
**Table 5.1. Illustration of Gay-Lussac's Law of Combining Volumes**

Reaction			Products	Volume Ratio
1. Hydrogen (Gas) + Chloride (Gas)	1 volume	1 volume	Hydrogen chloride (Gas) 2HCl	1 : 1 : 2
H <sub>2</sub> Cl <sub>2</sub>				
2. Nitrogen (Gas) + Oxygen (Gas)	1 volume	1 volume	Nitrogen(II) oxide (Gas) 2NO	1 : 1 : 2
N <sub>2</sub> O <sub>2</sub>				
3. Hydrogen (Gas) + Oxygen (Gas)	2 volume	1 volume	Steam (Gas) 2H <sub>2</sub> O	2 : 1 : 2
2H <sub>2</sub> O <sub>2</sub>				
4. Nitrogen (Gas)+ Hydrogen (Gas)	1 volume	3 volumes	Ammonia (Gas) 2NH <sub>3</sub>	1 : 3 : 2
N <sub>2</sub> 3H <sub>2</sub>				
5. Carbon monoxide (Gas) + Oxygen (Gas)	2 volumes	1 volumes	Carbon dioxide (Gas) 2CO <sub>2</sub>	2 : 1 : 2
2CO                            O <sub>2</sub>				
6. Hydrogen Sulphide (Gas) + Chlorine (Gas)	1 volume	1 volume	Hydrogen chloride (Gas) + Sulphur(Solid) 2HCl + S	1 : 1 : 2
H <sub>2</sub> S                            Cl <sub>2</sub>				

### Numerical problems based on Gay-Lussac's law of combining volumes

**Q. 1. What volume of oxygen would be required to burn 200 ml of acetylene ( $\text{C}_2\text{H}_2$ ) completely and what would be the volume of carbon dioxide formed ?**

**Solution :**



According to the equation, 2 volumes of acetylene require 5 volumes of oxygen for complete combustion to produce 4 volumes of carbon dioxide.

∴ 200 ml of acetylene will require

$$= \frac{5}{2} \times 200 \text{ ml} = 500 \text{ ml}$$

of oxygen, i.e., 500 ml of oxygen is required to produce  $200 \times 2 = 400$  ml of carbon dioxide.

Gay-Lussac's law states that when gases chemically react, they do so in volumes which bear a simple whole number ratio to each other and to the volumes of the products provided the products are also in a gaseous state under similar conditions of temperature and pressure.

### Table 5.1. Illustration of Gay-Lussac's Law of Combining Volumes

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N <sub>2</sub> O <sub>2</sub>				
3. Hydrogen (Gas) + Oxygen (Gas)	2 volume	1 volume	Steam (Gas) 2H <sub>2</sub> O	2 : 1 : 2
2H <sub>2</sub> O <sub>2</sub>				
4. Nitrogen (Gas)+ Hydrogen (Gas)	1 volume	3 volumes	Ammonia (Gas) 2NH <sub>3</sub>	1 : 3 : 2
N <sub>2</sub> 3H <sub>2</sub>				
5. Carbon monoxide (Gas) + Oxygen (Gas)	2 volumes	1 volumes	Carbon dioxide (Gas) 2CO <sub>2</sub>	2 : 1 : 2
2CO                            O <sub>2</sub>				
6. Hydrogen Sulphide (Gas) + Chlorine (Gas)	1 volume	1 volume	Hydrogen chloride (Gas) + Sulphur(Solid) 2HCl + S	1 : 1 : 2
H <sub>2</sub> S                            Cl <sub>2</sub>				

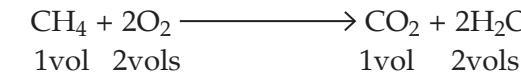
**Answer :** 500 ml of oxygen is required to burn 200 ml of acetylene completely, thereby producing 400 ml of carbon dioxide.

**Q. 2. Methane burns in oxygen to form carbon dioxide and water vapour as given by the equation :**



Calculate (i) the volume of oxygen needed to burn 50 cm<sup>3</sup> of methane completely and (ii) the volume of carbon dioxide formed.

**Solution :**



Since 1 volume of methane reacts with 2 volumes of oxygen,

∴ 50 cm<sup>3</sup> of methane will react with  $2 \times 50 = 100$  cm<sup>3</sup> of oxygen.

Similarly, 50 cm<sup>3</sup> of carbon dioxide is formed with 100 cm<sup>3</sup> of water.



∴ The volume of oxygen used = 100 cm<sup>3</sup>

∴ The volume of carbon dioxide formed = 50 cm<sup>3</sup>

**Q. 3. If 6 litres of hydrogen and 5.6 litres of chlorine are mixed and exploded, what will be the composition by volume of the resulting gaseous mixture?**

**Solution :**



Since 1 volume of chlorine reacts with 1 volume of hydrogen,

5.6 litres of chlorine will react with only 5.6 litres of hydrogen.

∴ (6 – 5.6), i.e., 0.4 litres of hydrogen will remain unreacted.

Since the volume of HCl gas formed is twice that of chlorine used,

∴ Volume of HCl formed will be = 5.6 × 2 = 11.2 litres.

**Answer :** 11.2 litres of HCl and 0.4 litres of residual hydrogen.

**Q. 4. What volume of propane is burnt for every 100 cm<sup>3</sup> of oxygen in the reaction?**



(Gas volumes measured under the same conditions)

**Solution :** From the above reaction, it is clear that for every 5 volumes of oxygen, 1 volume of propane is burnt.

∴ Volume of propane burnt for every 100 cm<sup>3</sup> of oxygen

$$= \frac{1}{5} \times 100 = 20 \text{ cm}^3$$

**Answer :** 20 cm<sup>3</sup> of propane is burnt.

**Q. 5. 80 cm<sup>3</sup> of methane is mixed with 200 cm<sup>3</sup> of pure oxygen at room temperature and pressure. The mixture is then ignited and it burns as illustrated by the equation :**



Calculate the composition of the resulting mixture if it is cooled to initial room temperature and pressure.

**Solution :**



By Gay-Lussac's law,

(i) 1 vol of methane requires 2 vols of oxygen.

∴ 80 cm<sup>3</sup> of methane requires = 2 × 80 = 160 cm<sup>3</sup> oxygen

(ii) 1 vol of methane produces 1 vol of carbon dioxide.

∴ 80 cm<sup>3</sup> of methane produces 80 cm<sup>3</sup> of carbon dioxide.

Hence, the composition of gaseous mixture after reaction is :

$$\text{Methane} = (80 - 80) = 0$$

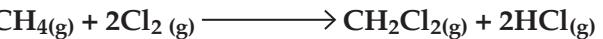
$$\text{Carbon dioxide} = 80 \text{ cm}^3$$

$$\text{Oxygen} = (200 - 160) = 40 \text{ cm}^3$$

Water = negligible.

**Q.6.40 cm<sup>3</sup> of methane reacts with chlorine according to the following equation.**

**Equation :**



Calculate the volume of HCl gas formed and chlorine gas required.

**Solution :**

By Gay-Lussac's law :

1 vol of CH<sub>4</sub> produces 2 vols of HCl.

∴ 40 cm<sup>3</sup> of CH<sub>4</sub> produces 2 × 40 = 80 cm<sup>3</sup> of HCl

Since 1 vol of CH<sub>4</sub> requires 2 vols of chlorine, 40 cm<sup>3</sup> of methane will require 2 × 40 = 80 cm<sup>3</sup> of chlorine.

**Answer :**

80 cm<sup>3</sup> of HCl is formed.

80 cm<sup>3</sup> of chlorine is required.

**Q.7.200 cm<sup>3</sup> of ethylene [C<sub>2</sub>H<sub>4</sub>] is burnt in just sufficient air (containing 20% oxygen) to form carbon dioxide gas and steam. If all measurements are made at constant pressure and 100°C, find the composition of the resulting mixture.**

**Solution :** The reaction involved is given by :



By Gay-Lussac's law :

(i) 1 vol of ethylene requires oxygen = 3 vols

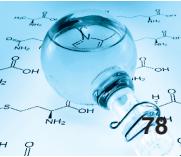
∴ 200 cm<sup>3</sup> of ethylene will require oxygen = 3 × 200 = 600 cm<sup>3</sup>.

(ii) 1 vol of ethylene produces carbon dioxide = 2 vols

∴ 200 cm<sup>3</sup> of ethylene will produce carbon dioxide = 2 × 200 = 400 cm<sup>3</sup>

(iii) 1 vol of ethylene produces steam = 2 vols

∴ 200 cm<sup>3</sup> of ethylene will produce steam = 2 × 200 = 400 cm<sup>3</sup>.



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When oxygen is 20%, unreacted air is = 80%.

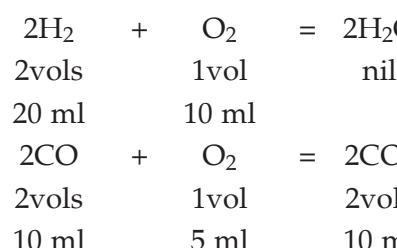
$$\text{When oxygen is } 600 \text{ cm}^3, \text{ then unreacted air is} \\ = \frac{80 \times 600}{20} = 2400 \text{ cm}^3$$

Hence, the composition of mixture after reaction :

- (i) Carbon dioxide = 400 cm<sup>3</sup>
- (ii) Steam = 400 cm<sup>3</sup>
- (iii) Unreacted air = 2400 cm<sup>3</sup>

**Q.8. 20 ml of hydrogen, 10 ml of carbon monoxide and 20 ml of oxygen are exploded in a eudiometer. What will be the volume and composition of the mixture of gases, after cooling to room temperature? Write each reaction separately.**

**Solution :**



Total volume of oxygen used = (10 + 5) = 15 ml.

Volume of oxygen left over = 20 - 15 = 5 ml.

Volume of carbon dioxide formed = 10 ml.

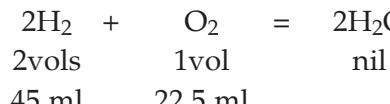
∴ Total volume of gases = 15 ml

**Answer :** The total volume of gaseous mixture after cooling to room temperature would be 15 ml, containing 5 ml oxygen and 10 ml carbon dioxide.

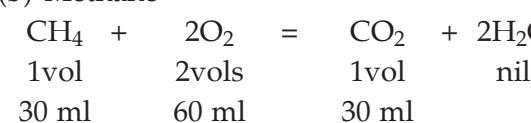
**Q.9. A sample of coal gas contained 45% H<sub>2</sub>, 30% CH<sub>4</sub>, 20% CO, and 5% C<sub>2</sub>H<sub>2</sub> by volume. 100 ml of this gaseous mixture was mixed with 160 ml of oxygen and exploded. Calculate the volume and the composition of the resulting mixture, when cooled to room temperature and pressure.**

**Solution :** 100 ml of gaseous mixture will contain 45 ml H<sub>2</sub>, 30 ml CH<sub>4</sub>, 20 ml CO and 5 ml C<sub>2</sub>H<sub>2</sub>. Equations representing the involved combustion are as follows :

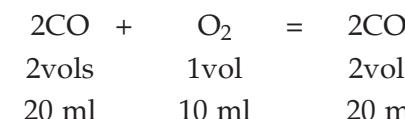
(a) Hydrogen



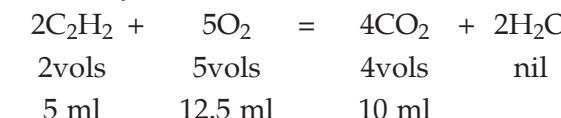
(b) Methane



(c) Carbon monoxide



(d) Acetylene



$$\begin{aligned} \text{Total volume of oxygen used} &= (22.5 + 60 + 10 + 12.5) \\ &= 105 \text{ ml} \end{aligned}$$

Volume of oxygen left = (160 - 105) = 55 ml

Total volume of carbon dioxide formed

$$= 30 + 20 + 10 = 60 \text{ ml}$$

∴ Total volume of gases left = (55 + 60) = 115 ml

**Answer :** 55 ml of residual oxygen and 60 ml of carbon dioxide were left. The total volume of gases left was 115 ml.

### 5.4 AVOGADRO'S LAW

Amedo Avogadro, in 1811, put forward a hypothesis based on the differentiation between the ultimate particle of an element that takes part in the reactions and the ultimate particle that exists independently. He made the following distinction between the atom and molecules of a gaseous element.

*The smallest particle of an element that takes part in a chemical reaction is called atom. It may or may not exist independently.*

*The smallest particle of a substance (element or a compound) that exists independently is called a molecule.*

Avogadro argued that since the smallest particle of a gas that exists independently is a molecule and not an atom; hence, the volume of gases should be related to the number of molecules and not the number of atoms. He put forward Avogadro's law, which states that

*"Equal volume of all gases under similar conditions of temperature and pressure contain the same number of molecules regardless of their chemical nature and physical properties".*

This means that if the temperature, pressure and volume of two gases are the same, then the two gases contain the same number of molecules, regardless of their identity. For example, 1 L of H<sub>2</sub> contains the same number of molecules as in 1 L of O<sub>2</sub>, or of any other gas, provided the volumes of all gases are measured under similar conditions of temperature and pressure.



Lorenzo Romano Amedeo Carlo Avogadro di Quaregna e di Cerreto, Count of Quaregna and Cerreto (9 August 1776, Turin, Piedmont-Sardinia – 9 July 1856), was an Italian scientist, most noted for his contribution to molecular theory now known as Avogadro's law. As a tribute to him, the number of elementary entities (atoms, molecules, ions or other particles) in 1 mole of a substance,  $6.022 \times 10^{23}$ , is known as the Avogadro constant, one of the seven SI base units and represented by NA.

### Atomicity

The number of atoms that a molecule of an element contains is called the atomicity of the element.

Depending on the number of atoms present, molecules are classified as :

**Table 5.2. Types of Atoms**

Types of Atom	Examples
<b>1. Monoatomic</b> A monoatomic molecule contains only one atom.	Inert gas like helium(He), neon(Ne), argon (Ar) etc.
<b>2. Diatomic</b> A diatomic molecule is composed of two similar atoms.	Hydrogen ( $H_2$ ), oxygen( $O_2$ ), chlorine( $Cl_2$ ), nitrogen( $N_2$ ), etc.
<b>3. Triatomic</b> A triatomic molecule is composed of three similar atoms.	Ozone gas( $O_3$ )
<b>4. Tetratomic</b> A tetratomic molecule is composed of four similar atoms.	Phosphorous ( $P_4$ )
<b>5. Octatomic</b> An octatomic molecule is composed of eight similar atoms.	Sulphur( $S_8$ )

Mole Day is celebrated on October 23 every year. Mole Day was proposed in an article in The Science Teacher in early 1980s. Inspired by the article, Maurice Oehler, a chemistry teacher (now retired) in Prairie du Chien, Wisconsin, created the National Mole Day Foundation in 1991.

### 5.5 RELATION BETWEEN VAPOUR DENSITY AND RELATIVE MOLECULAR MASS OF A GAS

#### Relative molecular mass

It is defined as the ratio of the mass of 1 molecule of the gas or vapour to the mass of 1 atom of hydrogen.

**Relative molecular mass of a gas =**

$$\frac{\text{mass of 1 molecule of the gas or vapour}}{\text{mass of 1 atom of hydrogen}}$$

#### Vapour density

It is defined as the ratio of the mass of a certain volume of the gas or vapour to the mass of the same volume of hydrogen at the same temperature and pressure.

**Vapour density (V.D.) =**

$$\frac{\text{mass of 1 volume of gas or vapour}}{\text{mass of 1 volume of hydrogen}}$$

Applying Avogadro's law,

**Vapour density (V.D.) =**

$$\frac{\text{mass of 1 volume of gas or vapour}}{\text{mass of 1 volume of hydrogen}}$$

Since hydrogen is diatomic,

$$\text{V.D.} = \frac{\text{mass of molecule of gas or vapour}}{2 \times \text{mass of 1 atom of hydrogen}}$$

Multiplying both sides by 2, we get

$$2 \times \text{V.D.} = \frac{\text{mass of 1 molecule of gas or vapour}}{\text{mass of 1 atom of hydrogen}}$$

Therefore,

$2 \times \text{V.D.}$  = relative molecular mass of the gas or vapour.

$2 \times \text{vapour density}$  = molecular weight.

#### Molecular Mass from Vapour Density

The density of a gas relative to that of hydrogen at the same temperature and pressure is known as the vapour density or the relative density of the gas.

Thus,

The vapour density of a gas =

$$\frac{\text{density of the gas at a given P and T}}{\text{density of hydrogen at the same P and T}}$$

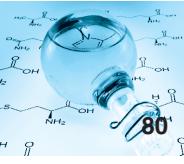
$$= \frac{\text{Mass of given Volume of the gas at given P and T}}{\text{Mass of the same volume of hydrogen at the same P and T}}$$

$$= \frac{\text{mass of 22.4 L of the gas at STP}}{\text{mass of 22.4 L of hydrogen at STP}}$$

$$= \frac{1 \text{ gram} - \text{molecular mass of the gas}}{2 \text{ g}}$$

∴ The relative molecular mass of a gas =  $\frac{1 \text{ gram} - \text{molecular mass of the gas}}{2 \text{ g}}$

The vapour densities of gases can be determined experimentally; so can be those of many other substances such as methanol, ethanol and acetone, which can be completely volatilized. Thus, it is easy to determine the relative molecular masses of such substances.



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**Table 5.3. Molecular Formula**

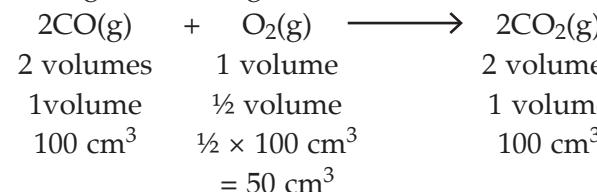
Substance	Vapour density	Relative molecular mass (= 2 × vapour density)	Molecular formula
Nitrogen	14	28	N <sub>2</sub>
Oxygen	16	32	O <sub>2</sub>
Chlorine	35.5	71	Cl <sub>2</sub>
Methane	8	16	CH <sub>4</sub>
Ethane	15	30	C <sub>2</sub> H <sub>6</sub>
Carbon dioxide	22	44	CO <sub>2</sub>
Methanol	16	32	CH <sub>3</sub> OH

### Problems Based on Avogadro's Law :

1. What is the volume of the oxygen required to burn 100 cm<sup>3</sup> of CO completely into CO<sub>2</sub>?

What is the volume of CO<sub>2</sub> produced? Assume that all volumes are measured at the same temperature and pressure.

**Solution :** The balance chemical equation representing the burning of CO into CO<sub>2</sub> is



So, the volume of oxygen required = 50 cm<sup>3</sup>

And that of carbon dioxide produced = 100 cm<sup>3</sup>

### 5.6. RELATIVE ATOMIC MASS (ATOMIC WEIGHT)

Atoms are too tiny and light to be seen or weighed directly to calculate their mass. One atom of hydrogen has a mass of  $1.66 \times 10^{-24}$  g; one atom of oxygen has a mass of  $1.99 \times 10^{-23}$  g. As these atomic masses are too small to handle, it is convenient to use the mass of some standard light atom to which the mass of other atoms is compared. The resultant masses thus obtained are Relative Atomic Masses (RAM) or atomic weight.

### 5.7. THE CARBON-12 SCALE

The standard with which the mass of other atoms is compared is the mass of one atom of C-12.

On this scale, one-twelfth the mass of an atom of the isotope <sup>12</sup>C is treated as the atomic mass unit (a.m.u), and relative atomic masses are determined accordingly.

Relative atomic mass is defined as follows :

The relative atomic mass of an element is the ratio of the mass of an atom of the element to one-twelfth the mass of an atom of carbon-12, or it is a number that shows how many times an atom of an element is heavier than one-twelfth the mass of an atom of carbon-12.

*The relative atomic mass of an element is a number that shows how many times an atom of an element is heavier than an atom of a reference element.*

**Relative atomic mass =**

$$\frac{\text{mass of 1 atom of the element}}{\frac{1}{12} \times \text{mass of 1 atom of } ^{12}\text{C}}$$

The relative atomic mass is expressed in atomic mass unit (a.m.u.), and it is defined as  $1/12^{\text{th}}$  the mass of one atom of C-12, i.e.,

$$1 \text{ a.m.u} = 1.66 \times 10^{-24} \text{ g}$$

Therefore, the mass of 1 atom of an element = relative atomic mass  $\times 1.66 \times 10^{-24}$  g

### Table 5.4. Atomic number and relative atomic mass of some important elements

Element	Symbol	Atomic number	Relative atomic mass (a.m.u.)
Hydrogen	H	1	1.0079
Helium	He	2	4.0026
Carbon	C	6	12.011
Nitrogen	N	7	14.007
Oxygen	O	8	15.999
Neon	Ne	10	20.180
Sodium	Na	11	22.990

### 5.8. FRACTIONAL ATOMIC MASSES

Most of the elements in nature contain two or more mixtures of isotopes. The relative atomic mass of an element is a number which shows how many times an atom of the element is heavier than an atom of a reference element.

Depending upon the reference element, different scales of atomic mass have been introduced from time to time. All of them have contributed significantly to the development of chemistry.

For example, chlorine consists of a mixture of two isotopes of masses 35 and 37 in the ratio of 3 : 1.

The average relative atomic mass of chlorine =  $\frac{35 \times 3 + 37 \times 1}{4} = 35.5$



## 5.9. RELATIVE MOLECULAR MASS (RMM) / (MOLECULAR WEIGHT)

Relative molecular mass is a number that indicates how many times a molecule of a substance is heavier in comparison to  $1/12^{\text{th}}$  of the mass of one carbon-12 atom. It is defined as :

*The ratio of the mass of a molecule of the substance to one-twelfth the mass of an atom of carbon-12.*

**Relative Molecular Mass =**

$$\frac{\text{mass of 1 molecule of the substance}}{\frac{1}{12} \times \text{mass of 1 atom of } 12\text{C}}$$

The relative molecular mass of a substance (element or compound) can be easily calculated by adding the relative atomic masses of all the individual atoms present in the molecule.

The unit of molecular mass is the same as that of atomic mass i.e. a.m.u.

**Example 1:** Calculate the molecular masses of the following compounds :

(a) Sucrose ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ )

(b) Sodium Hydroxide ( $\text{NaOH}$ )

(c) Sulphuric Acid ( $\text{H}_2\text{SO}_4$ )

(d) Ammonium Sulphate ( $(\text{NH}_4)_2\text{SO}_4$ )

1. Relative molecular mass of  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$

$$\begin{aligned} &= 12 \times 12 + 1 \times 22 + 16 \times 11 \\ &= 144 + 22 + 176 \\ &= 342 \text{ a.m.u.} \end{aligned}$$

2. Relative molecular mass of  $\text{NaOH}$

$$\begin{aligned} &= 23 + 16 + 1 \\ &= 40 \text{ a.m.u.} \end{aligned}$$

3. Relative molecular mass of  $\text{H}_2\text{SO}_4$

$$\begin{aligned} &= 2 \times 1 + 32 + 4 \times 16 \\ &= 98 \text{ a.m.u.} \end{aligned}$$

4. Relative molecular mass of  $(\text{NH}_4)_2\text{SO}_4$

$$\begin{aligned} &= 2(14+1\times 4) + 32 + (16\times 4) \\ &= 2\times 18 + 32 + 64 \\ &= 36 + 32 + 64 \\ &= 132 \text{ a.m.u.} \end{aligned}$$

## 5.10. GRAM ATOMIC MASS (GAM)/GRAM-ATOM :

The atomic mass of an element expressed in grams is called gram atomic mass or gram atom.

The quantity of the element that weighs equal to its gram atomic mass is called one-gram atom of that element.

For example, 1 gram atomic mass of oxygen, O (atomic weight approximately 16), is 16 grams.

**Example 1 :** How many gram-atoms are there in 80.0 g of oxygen?

**Solution :** The relative atomic mass of oxygen = 16

∴ The gram atomic mass of oxygen = 16 g

Given mass of oxygen = 80 g

∴ The number of gram-atoms =  $80 \text{ g}/16 \text{ g} = 5$

**Example 2 :** 2.5 gram-atoms of carbon are required for a particular reaction. How many grams of carbon would you use?

**Solution :** The relative atomic mass of carbon = 12

∴ 1 gram atomic mass of carbon = 12 g

∴ 2.5 gram atoms of carbon =  $2.5 \times 12 \text{ g} = 30 \text{ g}$ .

So, the amount of carbon to be used = 30g.

## 5.11. GRAM MOLECULAR MASS (GMM)

Gram molecular mass is defined as that much quantity of the substance (element or compound) whose mass in grams is numerically equal to its molecular mass. It is also one gram-molecule of the substance. For example, molecular mass of ammonia is 17 a.m.u.; therefore,

1 gram-molecule of  $\text{NH}_3$  = gram molecular mass of  $\text{NH}_3$  = 17 grams.

Likewise, 1 gram-molecule of  $\text{H}_2\text{O}$  = gram molecular mass of  $\text{H}_2\text{O}$  = 18 grams.

The unit of gram molecular mass of a substance is expressed in grams.

## 5.12. MOLE CONCEPT

To understand the amount of substances in a chemical change, we must know the number of particles, atoms or molecules of different substances in the reaction. That's not an easy task because such numbers are extremely large; therefore, to describe the number of particles, chemists use a quantity called mole.

A mole is defined as the amount of any substance that contains as many elementary entities (e.g. atoms, molecules, ions, electrons) as there are atoms in exactly 12 grams of pure C-12 isotope.

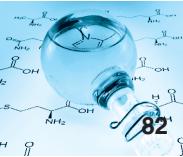
Just as a dozen eggs is a way of referring to exactly 12 eggs, a mole is a way of referring to  $6.022 \times 10^{23}$  molecules. This number of elementary entities is very large and is referred to as "Avogadro's Number".

Avogadro's number is defined as the number of atoms present in exactly 12 grams of C-12 isotope.

OR

Avogadro's number is the number of elementary units, i.e., atoms, ions or molecules present in one mole of a substance.

It is denoted by  $N_A$  or L and has a value equal to  $6.022 \times 10^{23}$ .



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It is necessary to specify the kind of particles involved, for which the mole concept is used. However, if the kind of particle is not expressed, then one mole is taken as  $6.022 \times 10^{23}$  molecules. For example, one mole of chlorine atoms represents  $6.022 \times 10^{23}$  atoms whereas 1 mole of chlorine represents  $6.022 \times 10^{23}$  molecules of  $\text{Cl}_2$ .

Likewise,

- One mole of oxygen atoms =  $6.022 \times 10^{23}$  atoms
- One mole of oxygen molecules =  $6.022 \times 10^{23}$  oxygen molecule
- One mole of ammonia molecules =  $6.022 \times 10^{23}$   $\text{NH}_3$  molecules
- One mole of chloride ions =  $6.022 \times 10^{23}$  chloride ions

### Mole of Atoms

"One mole of atoms contains  $6.022 \times 10^{23}$  atoms having mass equal to its gram atomic mass".

Thus, one mole of oxygen atoms contains  $6.022 \times 10^{23}$  atoms of oxygen and weight 16 g.

### Mole of Molecules

One mole contains  $6.022 \times 10^{23}$  molecules and is equivalent to the gram molecular mass of any given substance. Thus, 1 mole of  $\text{O}_2$  contains  $6.022 \times 10^{23}$  molecules and weighs 32g.

1 mole of water contains  $6.022 \times 10^{23}$  molecules and weighs 18g.

### Mole and Molar Volume

It has been noticed that one mole of any gaseous molecule occupies  $22.4 \text{ dm}^3$  (litre) or  $22,400 \text{ cm}^3$  (ml) at S.T.P. This volume is known as molar volume.

The molar volume of a gas can be defined as the volume occupied by one mole of a gas at S.T.P.

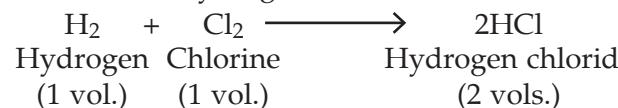
## 5.13 APPLICATIONS OF AVOGADRO'S HYPOTHESIS

### Atomicity of Gases

*The number of atoms that a molecule of an element contains is called the atomicity of the element.*

### Atomicity of hydrogen

In the reaction between hydrogen and chlorine, 1 vol. of hydrogen combines with 1 vol. of chlorine to form 2 vols. of hydrogen chloride.



Applying Avogadro's law,

$$1 \text{ molecule} + 1 \text{ molecule} = 2 \text{ molecules}$$

From the above equation, we can say that each molecule of hydrogen chloride must contain at least 1 atom of hydrogen.

Two molecules of hydrogen chloride, therefore, contain 2 atoms of hydrogen. These 2 atoms of hydrogen must have come from 1 molecule of hydrogen. Therefore, the molecule of hydrogen is diatomic and is written as  $\text{H}_2$ . Thus, the atomicity of hydrogen is 2.

Similarly, it can be shown that a molecule of chlorine contains 2 atoms, i.e., the atomicity of chlorine is 2. From the experimental evidence and similar arguments, it can be shown that oxygen and nitrogen also have 2 atoms per molecule, i.e., their molecules are diatomic.

The molecules of hydrogen, chlorine, oxygen and nitrogen are diatomic i.e., their atomicity is 2.

### Example :

1. The vapour density of  $\text{CO}_2$  is 22. What will be the volume of 88 g of carbon dioxide be at S.T.P?

### Solution :

The vapour density of  $\text{CO}_2$  = 22

$$\therefore \text{The relative molecular mass of } \text{CO}_2 \\ = 2 \times 22 = 44$$

$$\begin{aligned} \text{The number of moles of } \text{CO}_2 \text{ in 88 g of it} \\ = 88 \text{ g}/44 \text{ g} = 2 \end{aligned}$$

The volume of 1 mole of  $\text{CO}_2$  at S.T.P = 22.4 L

$$\begin{aligned} \text{Therefore, volume of 2 moles of } \text{CO}_2 \text{ at S.T.P} \\ = 44.8 \text{ L} \end{aligned}$$

So, the volume of 88 g of  $\text{CO}_2$  at S.T.P = 44.8 L

### Important Formulae

Mole and gram atomic mass :

$$\text{One mole of atoms} = 6.022 \times 10^{23} \text{ atoms}$$

$$\begin{aligned} &= \text{gram atomic mass of an element} \\ &= 1 \text{ g atom of the element.} \end{aligned}$$

Mole and gram molecular mass :

One mole of molecules

$$= 6.022 \times 10^{23} \text{ molecules}$$

$$= \text{gram molecular mass}$$

$$= 1 \text{ g molecule of the compound.}$$



Mole in terms of volume :

One mole of a gas = 22.4 litres at S.T.P.

Moles of an element = (Mass of the element) / (Atomic mass or G.A.W.)

Moles of a compound = (Mass of the compound) / (Molecular mass or G.M.W.)

Mass of one atom = (atomic mass or G.A.W.) /  $(6.022 \times 10^{23})$

Mass of one molecule = (molecular mass or G.A.W.) /  $(6.022 \times 10^{23})$

Number of molecules = moles  $\times 6.022 \times 10^{23}$

Number of atoms = moles  $\times 6.022 \times 10^{23}$

#### 5.14. PERCENTAGE COMPOSITION

Percentage composition of a compound is the percentage by weight of each element present in it.

**Percentage of an element in a compound**

$$= \frac{\text{total weight of the element in one molecule}}{\text{gram molecular weight of the compound}}$$

**Steps involved in calculating the percentage composition**

1. The weight of the element in one molecule is calculated.
2. The molecular weight of the compound is calculated.
3. The percentage of each element in the given compound is calculated by using the formula.

**Example 1 :** Calculate the percentage of H and O in  $\text{H}_2\text{O}$  (Given that relative molecular mass of H = 1 and O = 16).

**Solution :** The relative molecular mass of  $\text{H}_2\text{O}$

$$= 2 \times 1 + 16 = 18$$

18 g of  $\text{H}_2\text{O}$  contains 2 g of H and 16 g of O.

$\therefore$  100 g of  $\text{H}_2\text{O}$  contains  $2 \times 100/18 = 11.11$  g of H and  $16 \times 100/18 = 88.89$  g of O

So, water contains 11.11% H and 88.89% O.

**Note :** The sum of the percentage of all elements of a compound must be 100. In this case :  $11.11 + 88.89 = 100$ .

**Example 2 :** A 10.0 g sample of water has a percent composition of 88.8% oxygen and 11.2% hydrogen. (a) What is the percent composition of a 6.67 g sample of water? (b) Calculate the number of grams of oxygen in a 6.67 g sample of water.

**Solution :** 88.8% O and 11.2% H. The percent composition does not depend on the sample size.

$$(a) \quad 6.67 \text{ g } \text{H}_2\text{O} \left( \frac{88.8 \text{ g O}}{100 \text{ g H}_2\text{O}} \right) = 5.92 \text{ g O}$$

**Example 3 :** Calculate the percent composition of each of the following (a)  $\text{C}_3\text{H}_6$  and (b)  $\text{C}_5\text{H}_{10}$ .

**Solution :**

$$(a) \quad 1\text{C} = 12 \text{ a.m.u.}$$

$$\therefore \quad 3\text{C} = 12 \times 3$$

$$= 36 \text{ a.m.u.}$$

$$\text{Total} = 42.0 \text{ a.m.u.}$$

The percentage carbon is found by dividing the mass of carbon in one molecule by the mass of the molecule and multiplying the quotient by 100%.

$$\% \text{ C} = \left( \frac{36.0 \text{ amu C}}{42.0 \text{ amu total}} \right) \times 100\% \\ = 85.7\% \text{ C}$$

$$\% \text{ H} = \left( \frac{6.0 \text{ amu H}}{42.0 \text{ amu total}} \right)$$

The two percentages add up to 100.0%

$$1\text{C} = 12 \text{ a.m.u.}$$

$$5\text{C} = 12 \times 5$$

$$= 60 \text{ a.m.u.}$$

$$\text{Total} = 70.0 \text{ a.m.u.}$$

$$(b) \quad 1\text{H} = 1 \text{ a.m.u.}$$

$$\therefore \quad 6\text{H} = 6 \times 1 = 6 \text{ a.m.u.}$$

$$1\text{H} = 1 \text{ a.m.u.}$$

$$10\text{ H} = 1 \times 10$$

$$= 10 \text{ a.m.u.}$$

$$\% \text{ C} = \left( \frac{60.0 \text{ amu C}}{70.0 \text{ amu total}} \right) = 100\% \\ = 85.7\% \text{ C}$$

$$\% \text{ H} = \left( \frac{10.0 \text{ a.m.u. H}}{70.0 \text{ a.m.u. total}} \right) \\ 100\% = 14.3\% \text{ H}$$

The percentages are the same as those in part (a). Since the ratio of atoms of carbon to atoms of hydrogen is same (1 : 2) in both compounds, the ratio of masses should also be same and their percent by mass will be same too. From another viewpoint, this result also suggests that two compounds cannot be distinguished from each other by their percent compositions alone.

**Example 4 :** A certain mixture of salt ( $\text{NaCl}$ ) and sugar ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ ) contains 40% chlorine by mass. Calculate the percentage of salt in the mixture.

**Solution :** In 100 g of sample (the size does not make any difference), there is 40 g of chlorine, and therefore :

$$40 \text{ g Cl} \left( \frac{58.5 \text{ g NaCl}}{35.5 \text{ g Cl}} \right) = 65.9 \text{ g NaCl}$$



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The percentage of NaCl (in the 100g sample) is therefore 65.9%.

**Example 5 :** Calculate the percentage composition of Cu and water of crystallisation in blue vitriol ( $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ ).

**Solution :** The relative formula mass of  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ .

$$\begin{aligned} &= 63.5 + 32 + 4 \times 16 + 5(2 + 16) \\ &= 249.5 \end{aligned}$$

249.5 g of  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  contains 63.5 g of Cu and 90 g of  $\text{H}_2\text{O}$ .

$$\begin{aligned} \therefore 100 \text{ g of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O} &\text{ contains } 63.5 \times 100/249.5 \\ &= 25.45 \text{ g of Cu and } 90 \times 100/249.5 = 36.07 \text{ g of H}_2\text{O}. \end{aligned}$$

So, the blue vitriol contains 25.45% Cu and 36.07%  $\text{H}_2\text{O}$  (water of crystallisation).

### 5.15. EMPIRICAL FORMULA OF A COMPOUND

The simplest whole number ratio of atoms of different elements present in a compound is called an empirical formula.

Empirical formulae give no information of the actual numbers of atoms of each element in the molecule of the compound.

For example, the empirical formula of glucose is  $\text{CH}_2\text{O}$ . It indicates that the ratio of C, H and O atoms in a molecule of glucose is 1 : 2 : 1; whereas, its actual formula is  $\text{C}_6\text{H}_{12}\text{O}_6$ .

Similarly, the empirical formula of ethylene ( $\text{C}_2\text{H}_4$ ) is  $\text{CH}_2$ . It indicates that the ratio of C and H atoms in a molecule of ethylene is 1 : 2.

**Empirical formula mass** is the sum of the atomic masses of various elements present in the empirical formula.

Thus, for hydrogen peroxide ( $\text{H}_2\text{O}_2$ ), the empirical formula is HO and its empirical formula mass is  $1 + 16 = 17$ .

Empirical formula of a compound can be calculated knowing :

- (i) The percentage composition of the compound
- (ii) Atomic mass of the elements forming the compound.

1. Empirical formula of glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ ) is  $\text{CH}_2\text{O}$ . It indicates that the ratio of C, H and O atoms in a molecule of glucose is 1 : 2 : 1.

2. A compound is found to contain 40% carbon, 6.67% hydrogen and 53.34% oxygen. Find out the empirical formula.

$$\text{C} = 12, \text{H} = 1, \text{O} = 16$$

The mass of the elements in 100 parts by mass of the compound

$$\begin{aligned} &= \text{C} : \text{H} : \text{O} \\ &= 40 : 6.76 : 53.34 \end{aligned}$$

To find the ratio of the number of atoms in a molecule of the compound, the ratio of the mass of each element should be divided by the corresponding atomic mass.

Therefore, the ratio of the number of atoms of a molecule of the compound.

$$\begin{aligned} &= \text{C} : \text{H} : \text{O} \\ &= 40/12 : 6.76/1 : 53.34/16 \\ &= 3.33 : 6.76 : 3.33 \end{aligned}$$

To obtain the whole number ratio of atoms, the ratios obtained from these percentage compositions are to be divided by the smallest number found in the ratio as to obtain integer values for all the atoms present.

The ratio of the number of atoms in a molecule of the compound in a whole number.

$$\begin{aligned} &= \text{C} : \text{H} : \text{O} \\ &= 3.33/3.33 : 6.76/3.33 : 3.33/3.33 \\ &= 1 : 2 : 1 \end{aligned}$$

Thus, the empirical formula of the compound is  $\text{CH}_2\text{O}$ .

### 5.16. MOLECULAR FORMULA OF A COMPOUND

The molecular formula of a compound gives the actual number of each atom of different elements present in the molecule of a compound.

For example, the molecular formula of blue vitriol is  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ . It signifies that the blue vitriol is made up of :

- One atom of copper.
- One atom of sulphur.
- Four atoms of oxygen.
- Five molecules of water of crystallisation.

The molecular formula is either the same as the empirical formula or a multiple of it. For example, the empirical formula of water ( $\text{H}_2\text{O}$ ) is same as the molecular formula; whereas, the molecular formula of ethylene ( $\text{C}_2\text{H}_4$ ) is twice the empirical formula of ethylene ( $\text{CH}_2$ ).

However, the molecular formula does not convey any structural information.

#### Calculating Empirical Formula :

The empirical formula of a compound can be calculated from its percentage composition in the following steps :



- Tabulate the percentage composition of the compound. This gives the mass of each element present in 100 g of the compound.
- Divide the percentage of each element by its relative atomic mass to obtain the number of moles of atoms of the elements present in 100 g of the compound.
- Divide the number of moles by a suitable common factor to obtain them in the simplest whole-number ratio.
- Deduce the empirical formula of the compound from the whole-number ratio of the atoms.

#### Molecular formula :

Molecular formula can be derived from the empirical formula when the molecular mass of the compound is known.

Molecular formula = empirical formula  $\times$  n.  
where, n is an integer.

Therefore,

Relative molecular mass = n  $\times$  empirical formula mass.

#### Determination of Empirical and Molecular Formula

- The percentage composition of each element in a compound is determined by elemental analysis.
- The percentage composition of each element in a compound is divided by the corresponding atomic mass. The resulting number is called atomic ratio.
- The numbers obtained in step (ii) are then divided by a common factor (smallest number) when the simplest ratio of the number of atoms is obtained.
- From the ratio obtained in step (iii), the simplest whole number ratio is obtained by omitting small fractions or by multiplying the whole ratio by a common factor. The simplest whole number ratio gives the empirical formula.
- Then, the empirical formula mass is calculated.
- The molecular mass will be given or arrived at by using the relationship

Vapour density  $\times$  2 = molecular mass.

- The molecular mass of the compound is divided by the empirical formula mass when the value of 'n' is obtained.
- The molecular formula is arrived at when the empirical formula is multiplied by the value of n.

Molecular formula = empirical formula  $\times$  n  
Where, n = molecular weight / empirical formula weight.

**Table 5.6 : Difference between Empirical and Molecular Formula of a compound**

Empirical formula	Molecular formula
It is the simple formula.	It is the actual formula.
It corresponds to the percentage composition of a compound.	It corresponds to the molecular mass of a compound.
It represents the relative number of atoms of each element in one molecule of a compound.	It shows the actual number of each atom of the element in one molecule of a compound.

- Calculate the empirical formula of a compound if it contains 5.9% H and 94.1% O.

Element	%	RAM	Moles of atom	Atomic ratio
H	5.9	1	$\frac{5.9}{1} = 5.9$	$\frac{5.9}{5.9} = 1$
O	94.1	16	$\frac{94.1}{16} = 5.9$	$\frac{5.9}{5.9} = 1$

The empirical formula is HO

- Ammonium chloride contains 31.8% NH<sub>3</sub> and 66.3% Cl<sup>-</sup>. Work out the empirical formula of the compound (H=1, N = 14, O = 16, Cl = 35.5)  
NH<sub>3</sub> and Cl<sup>-</sup> account for 31.8 + 66.3 = 98.1% composition. The remaining percentage i.e. 100 - 98.1 = 1.9, must be taken for by the H<sup>+</sup> ion attached to NH<sub>3</sub>, forming NH<sub>4</sub><sup>+</sup>.

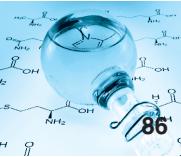
So, the percentage of NH<sub>4</sub><sup>+</sup> = 31.8 + 1.9 = 33.7

Ion	%	Relative ionic mass	Mole of ions	Ionic ratio
NH <sub>4</sub> <sup>+</sup>	33.7	18	$\frac{33.7}{18} = 1.87$	= 1
Cl <sup>-</sup>	66.3	35.5	$\frac{66.3}{35.5} = 1.87$	$\frac{1.87}{1.87} = 1$

Therefore, the empirical formula of ammonium chloride is NH<sub>4</sub>Cl.

Calculation of molecular mass :

- Deduce the molecular formula of a compound whose empirical formula is CH<sub>3</sub> and relative molecular mass 30.



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Let the molecular formula be  $\text{CH}_3$ .

The empirical formula mass =  $12 + 3 = 15$ .

We know that,

$$\begin{aligned} N \times 15 &= 30 \\ N &= \frac{30}{15} \end{aligned}$$

Therefore, the molecular formula of the compound is  $(\text{CH}_3)_2$  i.e.,  $\text{C}_2\text{H}_6$ .

### Calculation of Molecular Formula

**Example 1 :** The molecular formula of an organic compound is  $\text{CH}_3\text{COOH}$ . What is its empirical formula?

**Solution :** Molecular formula =  $\text{CH}_3\text{COOH}$ , i.e.,  $\text{C}_2\text{H}_4\text{O}_2$

∴ Ratio of C, H and O is  $2 : 4 : 2$

Simple ratio is, i.e.,  $1 : 2 : 1$

Empirical formula =  $\text{CH}_2\text{O}$

**Example 2 :** An organic compound contains 71.14% carbon, 14.86% hydrogen and 14% of nitrogen. Calculate the empirical formula.

Element	% composition	Ratio of atoms / molecule	Simplest ratio
		% composition atomic mass	previous column smallest number
C	71.14	$\frac{71.14}{12} = 5.93$	$\frac{5.93}{1} = 5.93$
H	14.86	$\frac{14.86}{1} = 14.86$	$\frac{14.86}{1} = 14.86$
N	14	$\frac{14}{14} = 1$	$\frac{1}{1} = 1$

The empirical formula is  $\text{C}_6\text{H}_{15}\text{N}$ .

**Example 3 :** An organic compound contains carbon, hydrogen, oxygen, nitrogen and chlorine. Find its empirical formula from the following data :

$\text{C} = 56.71\%$ ,  $\text{H} = 4.72\%$ ,  $\text{N} = 8.27\%$  and  $0.165\text{ g}$  of it gave  $0.1397\text{ g}$  of silver chloride. If the vapour density is 85, calculate the molecular formula.

**Solution :** Calculation of percentage of chlorine.

$143.5\text{ g}$  of silver chloride contains  $35.5\text{ g}$   $\text{Cl}_2$ .

$0.1397\text{ g AgCl}$  contains =  $35.5 \times 0.1397 / 143.5\text{ g Cl}_2$

Percentage of  $\text{Cl}_2$  =  $20.95\%$

Percentage of oxygen

$$= 100 - (56.71 + 4.72 + 8.27 + 20.95)$$

$$= 100 - 90.65$$

$$= 9.35\%$$

Element	% composition	Ratio of atoms / molecule	Simplest ratio
		% composition atomic mass	previous column smallest number
C	56.71	$\frac{56.71}{12} = 4.73$	$\frac{4.73}{0.58} = 8$
H	4.72	$\frac{4.72}{1} = 4.72$	$\frac{4.72}{0.58} = 8$
N	8.27	$\frac{8.27}{14} = 0.59$	$\frac{0.59}{0.58} = 1$
Cl	20.95	$\frac{20.95}{35.5} = 0.59$	$\frac{0.59}{0.58} = 1$
O	9.35	$\frac{9.35}{16} = 0.58$	$\frac{0.59}{0.58} = 1$

The empirical formula =  $\text{C}_8\text{H}_8\text{ONCl}$

The empirical formula mass,

$$\begin{aligned} &= (8 \times 12) + (8 \times 1) + (1 \times 16) + \\ &\quad (1 \times 14) + (1 \times 35.5) \\ &= 96 + 8 + 16 + 14 + 35.5 \\ &= 169.5 \end{aligned}$$

Molecular mass =  $2 \times$  vapour density

$$\begin{aligned} &= 2 \times 85 \\ &= 170 \end{aligned}$$

$n$  = molecular mass / empirical formula mass

$$\begin{aligned} &= 170 / 169.5 \\ &= 1 \end{aligned}$$

Molecular formula = (empirical formula)  $n$

$$= \text{C}_8\text{H}_8\text{ONCl}.$$

### 5.17. CHEMICAL EQUATIONS

A chemical equation is a concise form of representing a chemical change using symbols and formulae and signs.

The substances necessary for the reaction to take place are referred to as reactants.

The substances that are produced or formed during the chemical change are known as products.

**Chemical equation gives information about :**

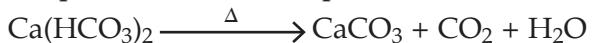
(I) Moles and masses of various reactants and products



(II) Volumes of gaseous reactants and products measured at S.T.P.

### Information Conveyed by the Chemical Equation

Examples for chemical equations :



1. ' $\Delta$ ' symbol represents heating. i.e., heat must be supplied in order to proceed with the reaction in forward direction.
2. The downward arrow ' $\downarrow$ ' indicates that the product is in the solid state and is precipitated.
3. The upward arrow ' $\uparrow$ ' denotes that the product is in the gaseous or vapour state.
4. The relative number of molecules represents the composition of the reactants and products.
5. The relative weights of reactants and products.
6. Relative volumes of gaseous substances.

Thus, a chemical equation gives a lot of information about a reaction.

### The following steps will help you to solve various problems based on chemical equations :

1. Write a chemical equation to represent the reaction and balance it.
2. Write the number of moles of each reactant and product. These numbers are the same as the numbers of molecules or ions on both sides of the equation.
3. Convert the number of moles of each reactant and product into mass or volume (if gaseous) as convenient or required.
4. Calculate the mass or volume of one or more of the reactants and products on the basis of the data given.

### Problems Based on Chemical Equation :

#### Based on Reacting Weights

**Example 1 :** Calculate the mass of the water formed by the reaction between 11.2 L of H<sub>2</sub> and 5.6 L of O<sub>2</sub> at S.T.P.

**Solution :**

The balanced chemical equation for the reaction is

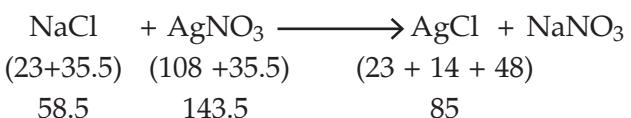
2H <sub>2</sub> (g)	+	O <sub>2</sub> (g)	$\longrightarrow$	2H <sub>2</sub> O (l)
2 mol		1 mol		2 mol
2 × 22.4 L (S.T.P.)	22.4 L (S.T.P.)			2 × 18 g
22.4 L (S.T.P.)	11.2 L (S.T.P.)			18 g
11.2 L (S.T.P.)	5.6 L (S.T.P.)			9 g

So, the mass of the water formed in the reaction = 9 g.

**Example 2 :** A solution of common salt when added to silver nitrate solution yields a precipitate

of silver chloride (0.28 g). Find the mass of sodium chloride in the solution and also the mass of sodium nitrate formed.

**Solution :**



Since, 85 g of silver chloride is formed from 58.5 g of NaCl.

Therefore,

$$\begin{aligned} &= 58.5 \times 0.28 / 85 \\ &= 16.38 / 85 \\ &= 0.192 \text{ g of NaCl} \end{aligned}$$

Hence, 0.192 g of NaCl is present in the solution.

Also, 58.5 g of sodium chloride with silver nitrate forms 85 g of NaNO<sub>3</sub>.

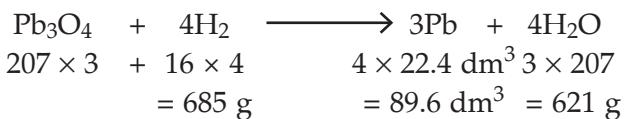
Therefore,

$$\begin{aligned} &= 85 \times 0.114 / 58.5 \\ &= 9.69 / 58.5 \\ &= 0.165 \text{ g of NaNO}_3 \end{aligned}$$

**Result :** 0.114 g of NaCl is present in the solution and 0.165 g of NaNO<sub>3</sub> is formed.

**Example 3 :** Find the mass of lead formed by the reduction of 342.5 g of red lead (Pb<sub>3</sub>O<sub>4</sub>) in a current of hydrogen and also find the volume of hydrogen used up at N.T.P.

**Solution :**



685 g of red lead on being reduced yields 621 g of lead.

Therefore,

$$\begin{aligned} &= \frac{621 \times 342.5}{685} \\ &= 310.5 \text{ g of lead} \end{aligned}$$

Also, 685 g of red lead uses 89.6 dm<sup>3</sup> of hydrogen at NTP.

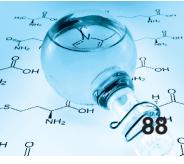
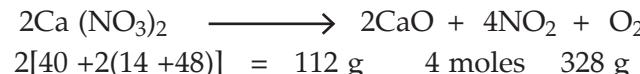
Therefore,

$$\begin{aligned} &= \frac{89.6 \times 342.5}{685} \\ &= 44.8 \text{ dm}^3 \text{ of hydrogen.} \end{aligned}$$

**Result :** 310.5 g of lead is formed and 44.8 dm<sup>3</sup> of hydrogen at N.T.P is used up.

#### Based on Mass-Volume Relationship

**Example 4 :** If 16.4 grams of calcium nitrate is heated, calculate the volume of (a) nitrogen dioxide obtained at S.T.P and (b) the mass of calcium oxide obtained.

**Solution :**(a) 328 g of  $\text{Ca}(\text{NO}_3)_2$  liberate  $4 \times 22.4 \text{ LNO}_2$ 

$$\therefore 16.4 \text{ g of } \text{Ca}(\text{NO}_3)_2 \text{ will liberate } \frac{4 \times 22.4}{328} \times 16.4 \\ = 4.48 \text{ L NO}_2 \text{ at S.T.P}$$

(b) 328 g of Calcium nitrate gives 112 g of CaO.

$$\therefore 16.4 \text{ g will give } = \frac{112 \times 16.4}{328} \\ = 5.6 \text{ g of CaO.}$$

**Result :**

- (a) 4.48 litres of  $\text{NO}_2$ .  
 (b) 5.6 g of CaO.

**INTEXT QUESTIONS**

- State Gay-Lussac's law of combining volumes.
- State Avogadro's law of volume.
- Define atomicity of gas. State the atomicity of hydrogen, phosphorous and sulphur.
- Differentiate between  $\text{N}_2$  and 2N.
- Explain why the number of atoms in a certain volume of hydrogen is twice the number of atoms in the same volume of helium at the same temperature and pressure.
- What are the volumes of hydrogen and oxygen required to produce 5 L of steam under the same conditions of temperature and pressure?
- What will be the volume of air containing 20%  $\text{O}_2$  required to completely burn 5 L of ethane ( $\text{C}_2\text{H}_4$ ) to give  $\text{CO}_2$  and steam? Assume that all volumes are measured at the same temperature and pressure.
- What is the volume of air required to completely burn 1 L of CO? Assume that all volumes are measured at the same temperature and pressure and also that air contains one-fifth by volume of  $\text{O}_2$ .
- A monovalent element in the gaseous state reacts with  $\text{Br}_2$  vapour to form its bromide, which is also gaseous. If 3 L of the monovalent element is taken, what is the volume of the  $\text{Br}_2$  vapour required and that of the product formed if all volumes are measured at the same temperature and pressure?
- Calculate the number of molecules in 2.5 moles of water.
- Calculate the number of moles when the number of atoms/molecules is given.

**Conversion Rules to Change**

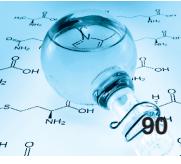
- Number of moles to gram : multiply the number of moles by molar mass.
- Grams to number of moles : divide the weight in grams by molar mass.
- Number of moles to number of particles: multiply the number of moles by Avogadro's number.
- Number of particles to number of moles: divide the number of particles by Avogadro's number.
- Number of moles to volume of a gas at S.T.P: multiply the number of moles by 22.4 L
- The volume of a gas at S.T.P to number of moles: divide the volume at S.T.P. by 22.4 L.

- (a) Convert the following into number of moles  
 (i)  $1.8 \times 10^{24}$  molecules of  $\text{CO}_2$  (ii)  $1.204 \times 10^{23}$  atoms of sodium.

- Calculate the number of moles in 9.2 g of sodium.
- Calculate the number of moles in 48 g of oxygen atoms.
- Calculate the number of moles in 6 g of hydrogen molecules, 11 g of carbon dioxide molecules, 20 g of sodium hydroxide and 24 g of water.
- The mass of one molecule of a substance is  $3.4 \times 10^{-22}$  g. Calculate its molecular mass.
- (a) The relative atomic mass of the Cl atom is 35.5 a.m.u. Explain this statement.  
 (b) What is the value of Avogadro's number?  
 (c) What is the value of the molar volume of a gas at S.T.P?
- Define or explain the terms : (a) Vapour density (b) molar volume (c) relative atomic mass (d) relative molecular mass (e) Avogadro's number (f) gram atom (g) mole.
- (a) What are the main applications of Avogadro's law ?  
 (b) How does Avogadro's law explain Gay-Lussac's law of combining volumes ?
- Calculate the relative molecular masses of  
 (a) ammonium chloroplatinate  $[(\text{NH}_4)_2 \text{PtCl}_6]$   
 (b) potassium chlorate (c)  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$   
 (d)  $(\text{NH}_4)_2\text{SO}_4$  (e)  $\text{CH}_3\text{COONa}$  (f)  $\text{CHCl}_3$   
 (g)  $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$



20. Find the (a) number of molecules present in 73 g of HCl (b) weight of 0.5 mole of O<sub>2</sub> (c) number of molecules in 1.8 g of H<sub>2</sub>O (d) number of moles in 10 g of CaCO<sub>3</sub> (e) weight of 0.2 mole of H<sub>2</sub> gas, (f) number of molecules in 3.2 g of SO<sub>2</sub>.
21. Which of the following would weigh most? (a) 1 mole of H<sub>2</sub>O (b) 1 mole of CO<sub>2</sub> (c) 1 mole of NH<sub>3</sub> (d) 1 mole of CO.
22. Which of the following contains maximum number of molecules? (a) 4g of O<sub>2</sub> (b) 4g of NH<sub>3</sub> (c) 4 g of CO<sub>2</sub> (d) 4 g of SO<sub>2</sub>
23. Calculate the number of (a) particles in 0.1mole of any substance (b) hydrogen atoms in 0.1 mole of H<sub>2</sub>SO<sub>4</sub>. (c) Molecules in 1 kg of CaCl<sub>2</sub>.
24. How many grams of (a) Al are present in 0.2 mole of it? (b) HCl are present in 0.1 mole of it? (c) H<sub>2</sub>O is present in 0.2 mole of it? (d) CO<sub>2</sub> is present in 0.1 mole of it ?
25. (a) The mass of 5.6 litres of a certain gas at S.T.P is 12 g. What is the relative molecular mass or molar mass of the gas? (b) Calculate the volume occupied at S.T.P by 2 moles of SO<sub>2</sub>.
26. Calculate the number of moles of (a) CO<sub>2</sub> which contain 8.00 g of O<sub>2</sub> (b) methane in 0.80 g of methane.
27. Calculate the actual mass of (a) an atom of oxygen (b) an atom of hydrogen (c) a molecule of NH<sub>3</sub> (d) an atom of silver (e) a molecule of oxygen (f) 0.25 gram atom of calcium.
28. Calculate the mass of 0.1 mole of each of the following (a) CaCO<sub>3</sub> (b) Na<sub>2</sub>SO<sub>4</sub>.10H<sub>2</sub>O (c) CaCl<sub>2</sub> (d) Mg
29. Calculate the number of oxygen atoms in 0.10 mole of Na<sub>2</sub>CO<sub>3</sub>.10H<sub>2</sub>O.
30. What mass of Ca will contain the same number of atoms as present in 3.2 g of S ?
31. Calculate the number of atoms in each of the following : (a) 52 moles of He (b) 52 a.m.u of He (c) 52 g of He.
32. Calculate the number of atoms of each kind in 5.3 grams of sodium carbonate.
33. (a) Calculate the mass of nitrogen supplied to soil by 5 kg of urea [CO (NH<sub>2</sub>)<sub>2</sub>]. (b) Calculate the volume occupied by 320 g of sulphur dioxide at STP.
34. (a) What do you understand by the statement 'vapour density of carbon dioxide is 22' ? (b) Atomic mass of chlorine is 35.5. What is its vapour density ?
35. What is the mass of 56 cm<sup>3</sup> of carbon monoxide at S.T.P?
36. Determine the number of molecules in a drop of water which weighs 0.09 g.
37. Give three kinds of information conveyed by the formula H<sub>2</sub>O.
38. Explain the terms empirical formula and molecular formula.
39. Give the empirical formula of C<sub>6</sub>H<sub>6</sub>, C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>, C<sub>2</sub>H<sub>2</sub>, CH<sub>3</sub>COOH<sub>2</sub>.
40. Find the percentage of water of crystallisation in the copper sulphate crystals, CuSO<sub>4</sub>.5H<sub>2</sub>O.
41. Calculate the percentage of phosphorus in Ca(H<sub>2</sub>PO<sub>4</sub>)<sub>2</sub>, Ca<sub>3</sub>(PO<sub>4</sub>)<sub>2</sub>.
42. Calculate the percentage composition of potassium chlorate, KClO<sub>3</sub>.
43. Find the empirical formula of the compounds with the following percentage composition. Pb = 62.5%, N = 8.5%, O = 29.0%.
44. Calculate the mass of iron in 10 kg of iron ore, which contains 80% of pure ferric oxide.
45. Acetylene and benzene have empirical formula CH and their vapour densities are 13 and 39, respectively. Find their molecular formula.
46. Find the empirical formula of a compound containing 17.7% hydrogen and 82.3% nitrogen.
47. On analysis, a substance was found to contain C = 54.54% H = 9.09% O = 36.36%. The vapour density of the substance is 44; Calculate (a) its empirical formula, and (b) its molecular formula.
48. An organic compound, whose vapour density is 45, has the following percentage composition. H = 2.22%; O = 71.19% and remaining carbon. Calculate (a) its empirical formula, (b) its molecular formula.
49. An organic compound contains 4.07% hydrogen, 71.65% chlorine and the remaining carbon. Its molar mass is 98.96. Find its (a) empirical formula and (b) molecular formula
14. A hydrocarbon contains 4.8 g of carbon per gram of hydrogen. Calculate (a) the g atom of each (b) find the empirical formula (c) find the molecular formula if its vapour density is 29.
50. Combine 0.2 g atoms of silicon with 21.3 g of chlorine. Find the empirical formula of the compound formed.
51. A hydrocarbon with molecular mass 78 a.m.u contains 7.7% by mass of hydrogen. Calculate the molecular formula of this compound.



52. In a compound of magnesium ( $Mg = 24$ ) and nitrogen ( $N = 14$ ), 18 g of magnesium combines with 7 g of nitrogen. Deduce the simplest formula by answering the following questions :
- How many gram-atoms of magnesium are equal to 18 g?
  - How many gram atoms of nitrogen are equal to 7 g of nitrogen?
  - Calculate the simple ratio of gram atoms of magnesium to gram-atoms of nitrogen and hence the simplest formula of the compound formed.
53. Barium chloride crystals contain 14.8% water of crystallisation. Find the number of molecules of water of crystallisation per molecule of barium chloride.
54. Urea is a very important nitrogenous fertilizer. Its formula is  $CON_2H_4$ . Calculate the percentage of nitrogen in urea.
55. Determine the formula of the organic compound if its molecule contains 12 atoms of carbon. The percentage compositions of hydrogen and oxygen are 6.48 and 51.42, respectively.
56. A compound with empirical formula  $AB_2$  has the vapour density equal to its empirical formula weight. Find its molecular formula.
- (b) A compound with empirical formula  $AB$  has vapour density three times its empirical formula weight. Find its molecular formula.
57. A hydride of nitrogen contains 87.5 percent by mass of nitrogen. Determine the empirical formula of this compound.
58. A compound has  $O = 61.32\%$ ,  $S = 11.15\%$ ,  $H = 4.88\%$  and  $Zn = 22.65\%$ . The relative molecular mass of the compound is 287 a.m.u. Find the molecular formula of the compound, assuming that all the hydrogen is present as water of crystallisation.



### POINTS TO REMEMBER

- ◆ Gay-Lussac's law of combining volumes: the volumes of the gases taking part in a chemical reaction and those of the products, if gaseous, are in a whole-number ratio under the same conditions of temperature and pressure.
- ◆ Avogadro's law : equal volumes of all gases, under the same conditions of temperature and pressure, contain an equal number of molecules.
- ◆ The number of atoms contained in a molecule of an element is called the atomicity of the element.
- ◆ The volumes of the reacting gases and the gaseous products, if any, under the same conditions of temperature and pressure, bear the same ratio as their co-efficient in a balanced chemical equation.
- ◆ The relative atomic mass of an element is a number that shows how many times an atom of the element is heavier than an atom of the reference element.
- ◆ On the hydrogen scale, the atomic mass of an element is the number of times an atom of the element is heavier than an atom of hydrogen.
- ◆ On the carbon-12 scale, one twelfth the mass of a carbon-12 atom is considered to be the atomic mass unit (a.m.u).
- ◆ The gram-atomic mass or the gram-atom or the mole of atoms of an element is the atomic mass of the element expressed in grams.
- ◆ The relative molecular mass of a substance is the ratio of the mass of a molecule of the substance to one twelfth the mass of an atom of carbon-12.
- ◆ The gram-molecular mass or the mole of a substance is the molecular mass of the substance expressed in grams.
- ◆ A mole is the amount of a substance that contains the same number of particles – atoms, molecules or ions – as the number of atoms in exactly 12 g of carbon-12.
- ◆ The volume occupied by 1 g molecular mass or 1 mole of a gas at a given temperature and pressure is called the molar volume of the gas.
- ◆ The standard molar volume of a gas is the molar volume of the gas at S.T.P. This volume is equal to 22.4L.
- ◆ Relative molecular mass =  $2 \times$  vapour density.
- ◆ The empirical formula of a compound shows the atomic ratio of the elements present in a molecule of the compound.
- ◆ The molecular formula of a compound gives the actual numbers of atoms of the different elements present in a molecule of the compound.
- ◆  $n \times$  empirical formula = relative molecular mass



## EXERCISES

## A. FILL IN THE BLANKS

- At the same temperature and pressure, equal volumes of gases contain the same number of .....
  - One mole of a gas at.....occupies ..... and is known as.....
  - There is ..... difference between the atomic and molecular masses of noble-gas elements
  - A compound of relative molecular mass 34 and empirical formula HO has the molecular formula .....
  - Which of the following would weigh the least ?
    - 2 gram atoms of Nitrogen
    - 1 mole of silver.
    - 22.4 litres of oxygen gas at 1 atmosphere pressure and 273 K.
    - $602 \times 10^{23}$  atoms of carbon [Atomic masses : Ag = 108, N = 14, O = 16, C = 12]
- [2015]

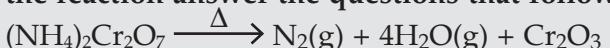
## B. MULTIPLE CHOICE TYPE QUESTIONS

Choose the Correct Answer from the options given below :

- The atomic mass of oxygen is 16 and the molecular mass of ozone is 48. What is the atomicity of ozone if it is an allotrope of oxygen
  - 1
  - 2
  - 3
  - 4
- The relative atomic mass of Ne is 20. Its relative molecular mass is :
  - 10
  - 20
  - 30
  - 40
- The volume occupied by half a mole of a gas at S.T.P is :
  - 5.6L
  - 11.2L
  - 22.4L
  - 44.8L
- Which of the following gases contains the largest number of molecules?
  - 40 g of hydrogen
  - 4 moles of SO<sub>2</sub>
  - 35.5 g of Cl<sub>2</sub>
  - 40 g of O<sub>2</sub>

## C. ANSWER THE FOLLOWING QUESTIONS BASED ON THE INSTRUCTIONS GIVEN BELOW

- (a) Consider the following reaction and based on the reaction answer the questions that follow :



Calculate :

- The quantity in moles of (NH<sub>4</sub>)<sub>2</sub>Cr<sub>2</sub>O<sub>7</sub> if 63 gm of (NH<sub>4</sub>)<sub>2</sub>Cr<sub>2</sub>O<sub>7</sub> is heated.

- The quantity in moles of nitrogen formed.

- The volume in litres or dm<sup>3</sup> of N<sub>2</sub> evolved at S. T. P.

- The mass in grams of Cr<sub>2</sub>O<sub>3</sub> formed at the same time.

[Atomic masses : H = 1, Cr = 52, N = 14]  
[2015]

## (b) Complete the following calculations

- Calculate the mass of Calcium that will contain the same number of atoms as are present in 3.2 gm of sulphur.  
(Atomic masses : S = 32, Ca = 40)
- If 6 litres of hydrogen and 4 litres of chlorine are mixed and exploded.

## D. MATCH THE COLUMNS

Column I	Column II
1. Boyle's law	Law of combining volumes
2. One mole of a gas at STP	$V_1 T_2 = V_2 T_1$
3. Molecular mass	$6.022 \times 10^{23}$ molecules
4. Gay lussac	( $\times$ Vapour density)
5. One mole	22.4 litres
6. Charle's law	$P_1 V_1 = P_2 V_2$

## E. NUMERICAL PROBLEMS

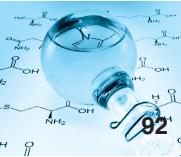
## Problems based on combining volumes

- When heated, potassium permanganate decomposes according to the following equation :  
 $2KMnO_4 \longrightarrow K_2MnO_4 + MnO_2 + O_2$

- Some potassium permanganate was heated in the test tube. After collecting one litre of oxygen at room temperature, it was found that the test tube had undergone a loss in mass of 1.32 g. If one litre of hydrogen under the same conditions of temperature and pressure has a mass of 0.0825 g; calculate the relative molecular mass of oxygen.

Given that the molecular mass of potassium permanganate is 158, what volume of oxygen (measured at room temperature) would be obtained by the complete decomposition of 15.8 g of potassium permanganate? (Molar volume at room temperature is 24 L)

- A flask contains 3.2 g of sulphur dioxide. Calculate the following: (i) the moles of sulphur



## CERTIFICATE CHEMISTRY-X

dioxide present in the flask (ii) the number of molecules of sulphur dioxide present in the flask (iii) The volume occupied by 3.2 g of sulphur dioxide at S.T.P

(b) An experiment showed that in a lead chloride solution, 6.21 g of lead is combined with 4.26 g of chlorine. What is the empirical formula of this chloride ?

3. The volumes of gases A, B, C and D are in the ratio, 1 : 2 : 2 : 4 under the same conditions of temperature and pressure.

(i) Which sample of gas contains the maximum number of molecules ?  
(ii) If the temperature and the pressure of gas A are kept constant, then what will happen to the volume of A when the number of molecules is doubled ?

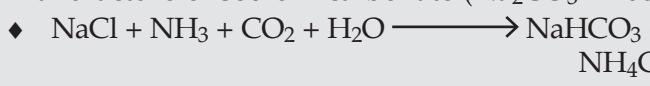
(iii) If this ratio of gas volume refers to the reactants and products of a reaction, which gas law is being observed ?

(iv) If the volume of A is actually 5.6 dm<sup>3</sup> at S.T.P, calculate the number of molecules in the actual volume of D at S.T.P. (Avogadro's number is  $6 \times 10^{23}$ )

(v) Using your answer from (iv) state the mass of D if the gas is dinitrogen oxide (N<sub>2</sub>O)

4. Calculate the percentage of nitrogen in aluminium nitride.

5. The equation given below relates to the manufacture of sodium carbonate (Na<sub>2</sub>CO<sub>3</sub> = 106)



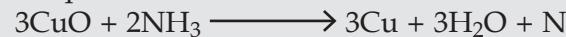
Questions (a) and (b) are based on the production of 21.2g of sodium carbonate.

- (a) What mass of sodium hydrogen carbonate must be heated to give 21.2 g of sodium carbonate ?  
(b) To produce the mass of sodium hydrogen carbonate calculated in (a) what volume of carbon dioxide, measured at S.T.P would be required ?

6. (a) Determine the empirical formula of a compound containing 47.9% potassium, 5.5% beryllium and 46.6% fluorine by mass.

(b) Given that the relative molecular mass of copper oxide is 80, what volume of ammonia (measured at S.T.P) is required to completely reduce 120 g of copper oxide ?

The equation for the reaction is :



7. (a) Calculate the number of moles and the number of molecules present in 1.4 g of ethylene gas.

What is the volume occupied by the same amount of ethylene ?

(b) What is the vapour density of ethylene?

8. (a) Calculate the percentage of sodium in sodium aluminium fluoride (Na<sub>3</sub>AlF<sub>6</sub>) correct to the nearest whole number.

(b) 560 ml of carbon monoxide is mixed with 500 ml of oxygen and ignited. The chemical equation for the reaction is as follows :



Calculate the volume of oxygen used and carbon dioxide formed in the above reaction.

9. A sample of ammonium nitrate when heated yields 8.96 litres of steam (measured at S.T.P)



(a) What volume of dinitrogen oxide is produced at the same time as 8.96 litres of steam?

(b) What mass of ammonium nitrate should be heated to produce 8.96 litres of steam? (Relative molecular mass of ammonium nitrate is 80.)

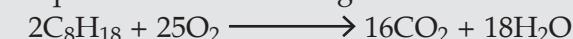
(c) Determine the percentage of oxygen in ammonium nitrate.

10. A compound X consists of 4.8% carbon and 95.2% bromine by mass.

(i) Determine the empirical formula of this compound working correct to one decimal place (C = 12; Br = 80).

(ii) If the vapour density of the compound is 252, what is the molecular formula of the compound ?

11. The equation for the burning of octane is :



(a) How many moles of carbon dioxide are produced when one mole of octane burns ?

(b) What volume, at S.T.P is occupied by the number of moles determined in (a) ?

(c) If the relative molecular mass of carbon dioxide is 44, what is the mass of carbon dioxide produced by burning two moles of octane ?

(d) What is the empirical formula of octane ?

12. (a) (i) A compound has the following percentage composition by mass: carbon 14.4%, hydrogen 1.2% and chlorine 84.5%. Determine the empirical formula of this compound. Work correct to 1 decimal place. (H = 1; C = 12; Cl = 35.5)

(ii) The relative molecular mass of this compound is 168, so what is its molecular formula ?

(b) From the equation



**Calculate :**

- (i) The mass of carbon oxidised by 49 g of sulphuric acid.  
(ii) The volume of sulphur dioxide measured at S.T.P liberated at the same time.

**F. VERY SHORT QUESTIONS**

- What is the gas equation, give its mathematical form ?
- State the Avogadro's law.
- State the Atomicity of the following
 

(a) Ne	(b) O <sub>3</sub>
(c) S <sub>8</sub>	(d) P <sub>4</sub>
- Give one application of Avogadro's hypothesis.
- Give the mathematical form of the following
 

(a) Boyle's law	(b) Charle's law
-----------------	------------------
- Give the Relationship between vapour density and relative molecular mass of a gas [formula only].
- Define a mole.
- Define the gram atomic mass of an atom.
- What is the term use for the ratio of the mass of a certain volume of gas to the mass of an equal

volume of hydrogen under the same conditions of temperature and pressure. [2014]

**G. SHORT QUESTIONS**

- State Gay-Lussac's law of combining volumes.
- What is Avogadro's law ?
- What do you mean by the atomicity of an element ?
- What is the relative atomic mass of an element ?
- Define relative molecular mass.
- What do you understand by the term 'molar volume of a gas' ?
- Define the standard molar volume of a gas.
- Define a mole. [ICSE 2009]
- What do you mean by the empirical formula of a compound ?

**H. LONG QUESTIONS**

- Explain Gay-Lussac's law of combining volumes on the basis of Avogadro's law.
- Briefly explain the applications of Avogadro's law.
- What is the significance of the number  $6.022 \times 10^{23}$  in chemistry ?

**RECENT BOARD QUESTIONS****Question 1.**

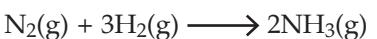
- (a) Choose the correct answer from the options given below :
- The vapour density of carbon dioxide [C = 12, O = 16]. [2012]
 

(A) 12	(B) 16
(C) 44	(D) 22
  - Which of the following would weigh the least ? [2015]
 

(A) 2 gram atoms of Nitrogen
(B) 1 mole of silver
(C) 22.4 litres of oxygen gas at 1 atmospheric pressure and 273 K
(D) $6.02 \times 10^{23}$ atoms of carbon [Atomic masses : Ag = 108, N = 14, O = 16, C = 12]

**Question 2.**

- (a) (i) 67.2 litres of hydrogen combines with 44.8 litres of nitrogen to form ammonia under specific conditions as : [2012]



Calculate the volume of ammonia produced. What is the other substance, if any, that remains in the resultant mixture ?

- The mass of  $5.6 \text{ dm}^3$  of a certain gas at STP is 12.0 g. Calculate the relative molecular mass of the gas.
- Find the total percentage of Magnesium in magnesium nitrate crystals, Mg(NO<sub>3</sub>)<sub>2</sub>.6H<sub>2</sub>O. [Mg = 24, N = 14; O = 16 and H = 1]

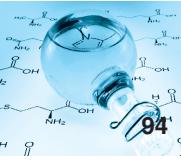
**Question 3.**

- (a) Solve the following : [2013]
- What volume of oxygen is required to burn completely  $90 \text{ dm}^3$  of butane under similar conditions of temperature and pressure ?  

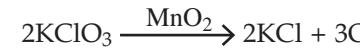
$$2\text{C}_4\text{H}_{10} + 13\text{O}_2 \longrightarrow 8\text{CO}_2 + 10\text{H}_2\text{O}$$
  - The vapour density of a gas is 8. What would be the volume occupied by 24.0 g of the gas at STP ?
  - A vessel contains X number of molecules of hydrogen gas at a certain temperature and pressure. How many molecules of nitrogen gas would be present in the same vessel under the same conditions of temperature and pressure ?

**Question 4.**

- O<sub>2</sub> is evolved by heating KClO<sub>3</sub> using MnO<sub>2</sub> as a catalyst : [2013]



## CERTIFICATE CHEMISTRY-X



- (i) Calculate the mass of  $\text{KClO}_3$  required to produce 6.72 litre of  $\text{O}_2$  at STP. [atomic masses of K = 39, Cl = 35.5, O = 16].
- (ii) Calculate the number of moles of oxygen present in the above volume and also the number of molecules.
- (iii) Calculate the volume occupied by 0.01 mole of  $\text{CO}_2$  at STP.

### Question 5.

- (a) (i) Oxygen oxidizes ethyne to carbon dioxide and water as shown by the equation : [2014]



What volume of ethyne gas at S.T.P. is required to produce 8.4  $\text{dm}^3$  of carbon dioxide at S.T.P.? [H = 1, C = 12, O = 16]

- (ii) A compound made up of two element X and Y has an empirical formula  $\text{X}_2\text{Y}$ . If the atomic weight of X is 10 and that of Y is 5 and the compound has a vapour density 25, find its molecular formula. [2014]

- (b) (i) State Avogadro's Law. [2014]

- (ii) A cylinder contains 68 g of ammonia gas at S.T.P. [2014]

1. What is the volume occupied by this gas ?
2. How many moles of ammonia are present in the cylinder ?
3. How many molecules of ammonia are present in the cylinder ? [N=14, H=1]

### Question 6.

- (a) Complete the following calculations. Show working for complete credit : [2015]

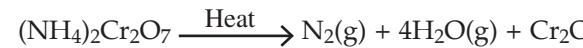
- (i) Calculate the mass of Calcium that will contain the same number of atoms as are present in 3.2 gm of Sulphur.

[Atomic masses : S = 32, Ca = 40]

- (ii) If 6 litres of hydrogen and 4 litres of chlorine are mixed and exploded and if water is added to the gases formed, find the volume of the residual gas.

- (iii) If the empirical formula of a compound is CH and it has a vapour density of 13, find the molecular formula of the compound

- (b) Consider the following reaction and based on the reaction answer the questions that follow : [2015]



### Calculate :

- (i) The quantity in moles of  $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$  if 63 gm of  $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$  is heated.
- (ii) The quantity in moles of nitrogen formed.
- (iii) The volume in litres or  $\text{dm}^3$  of  $\text{N}_2$  evolved at S. T. P.
- (iv) The mass in grams of  $\text{Cr}_2\text{O}_3$  formed at the same time.

[Atomic masses : H = 1, Cr = 52, N = 14]

### Question 7.

- (a) 1. A gas of mass 32 gms has a volume of 20 litres at S.T.P. Calculate the gram molecular weight of the gas. [2016]
2. How much calcium oxide is formed when 82g of calcium nitrate is heated ? Also find the volume of nitrogen dioxide evolved : [2016]



(Ca = 40, N = 14, O = 16)

- (b) A gas cylinder contains  $12 \times 10^{24}$  molecules of oxygen gas. If Avogadro's number is  $6 \times 10^{23}$ . Calculate : [2016]
1. The mass of oxygen present in the cylinder.
  2. The volume of oxygen at S.T.P. present in the cylinder. [O = 16]
- (c) A gaseous hydrocarbon contains 82.76% of carbon. Given that its vapour density is 29, find its molecular formula. [C = 12, H = 1] [2016]
- (d) The equation  $4\text{NH}_3 + 5\text{O}_2 \rightarrow 4\text{NO} + 6\text{H}_2\text{O}$ , represents the catalytic oxidation of ammonia. If 100  $\text{cm}^3$  of ammonia is used calculate the volume of oxygen required to oxidise the ammonia completely. [2016]

### Question 8.

- (a) (i) Calculate the volume of 320 g of  $\text{SO}_2$  at S.T.P (Atomic mass : S = 32 and O = 16). [2011]
- (ii) State Gay-Lussac's Law of combining volumes. [2011]
- (iii) Calculate the volume of oxygen required for the complete combustion of 8.8 g of propane ( $\text{C}_3\text{H}_8$ ). (Atomic mass : C = 14, O = 16, H = 1, Molar Volume = 22.4  $\text{dm}^3$  at STP). [2011]
- (b) Calculate the mass of
  - (i)  $10^{22}$  atoms of sulphur.
  - (ii) 0.1 mole of carbon dioxide.
- [Atomic mass : S = 32, C = 12 and O = 16 and Avogadro's Number =  $6 \times 10^{23}$ ]



# 6

# ELECTROLYSIS

## LEARNING OUTCOMES

- 6.0. Introduction
- 6.1. Common Terms in Cell
- 6.2. Oxidation and Reduction in Terms of Electron Transfers

- 6.3. Electrochemical Series of Metals
- 6.4. Discharge of Ions from Aqueous Solutions
- 6.5. Applications of Electrolysis
- 6.6. Electroplating
- 6.7. Electrefining

## 6.0 INTRODUCTION

Electrolysis is the process of separating of a compound into its constituent elements by passing an electrical current through it in its molten or aqueous state. The word *electrolysis* means the process of breaking molecules to smaller parts by using an electric current.

Some substances allow electricity to pass through them whereas some do not. All the elements that conduct electricity may not undergo chemical changes.

For example, elements like copper, aluminum, iron etc., pass electricity without undergoing chemical reaction and non-metals do not allow electricity to pass except graphite which conducts electricity.

In the electrolysis we have a electrolytic cell where the reaction is carried out, an electrolyte, which is a solution poured in the electrolytic cell and two electrodes which are connected by a battery.

### Metallic Conductors

A material which contains movable electric charges is called a conductor. Copper, aluminium, etc., are metallic conductors that possess charged movable electron particles. The current, consists of a single

stream of electrons flowing from the negative pole to the positive pole. Metals are good conductors of electricity both in solid and liquid state.

### Theory of Electrolytic Conductance:

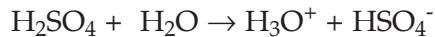
The Arrhenius theory of electrolytic conductance is also known as Arrhenius theory of ionization as it considers the electrolytic dissociation of ions.

### Postulates of Arrhenius Theory:

1. An electrolytic solution dissociates into free ions even if no current is passed through the solution. The positively charged particles are called cations and those having negative charges are called anions.
2. The ions present in solution constantly reunite to form neutral molecules. Thus, there is a state of dynamic equilibrium between the undissociated molecules and the ions, and the solution remains always electrically neutral since the total charge of the cations is always equal to the total charge of the anions.
3. The charged ions are free to move through the solution to the oppositely charged electrode. This is known as the migration of ions. This

movement of the ions is responsible for the flow of electric current through electrolytes. This explains the conductivity of electrolytes as well as the phenomenon of electrolysis.

- The electrical conductivity of an electrolytic solution depends on the number of ions present in the solution. Thus the degree of dissociation of an electrolyte determines whether it is a strong electrolyte or a weak electrolyte.
- The properties of solution of the electrolytes are the same as the properties of the ions in the solution of an electrolyte. As a whole the solution is electrically neutral unless an electric field is applied to the electrodes dipped into it. Presence of hydrogen ions ( $H^+$ ) renders the solution acidic while presence of hydroxide ions ( $OH^-$ ) renders the solution basic.
- There are two types of electrolytes – strong and weak. In a strong electrolyte, ionization is always incomplete because the ions impede each other's migration. This process is less in dilute solutions of strong electrolytes. Example, potassium chloride, sodium chloride and sulphuric acid.



In case of weak electrolytes, there is partial dissociation of ions in water and equilibrium exists between the dissociated ions and the undissociated electrolyte. In a weak electrolyte the degree of dissociation (ionization) is incomplete. Example : formic acid, acetic acid and ammonium hydroxide.



*Electrolysis was discovered by an English chemist William Nicholson who was the first person to produce chemical reaction through electricity. He broke water molecules to hydrogen and oxygen by placing wires from a battery into water.*

### Electrolytes

A compound that decomposes into ions on passage of electricity, either in aqueous solution or in molten state is called an electrolyte. Examples of electrolytes are sodium chloride solution, copper sulphate solution and acidified water. When electricity passes through an electrolyte, a chemical reaction takes place, which allows the decomposition of the electrolyte.

### Non-Electrolytes :

Some substances either in aqueous solution or in fused state do not conduct electricity. Such substances are called non-electrolytes. Examples are glucose, sugar, distilled water, urea.

## 6.1 COMMON TERMS IN CELL

### Current

It is the flow of electrons (electric charge) through a wire or any conductor.

### Electrodes

It is the material such as a metallic rod or bar or strip immersed in the electrolyte, which conducts electrons in to and out of a solution. Example: zinc, graphite, iron, aluminum, etc.

### Anode

Anode is the negatively charged electrode in an electric circuit. Positive current flows through anode in a closed circuit, and electrons leave the electrolyte through the anode.

It is connected to the positive terminal of the battery and is shown as (-) in the cell diagram.

### Cathode

It is the electrode at which electrons are received from the outer circuit. It has a positive charge and is shown as (+) in the cell diagrams. The cathode is connected to the negative terminal of the battery.

### Electrolytic Cell

It is the vessel where electrolysis is carried out during the process of electrolysis. The passed electrical energy is converted into chemical energy.

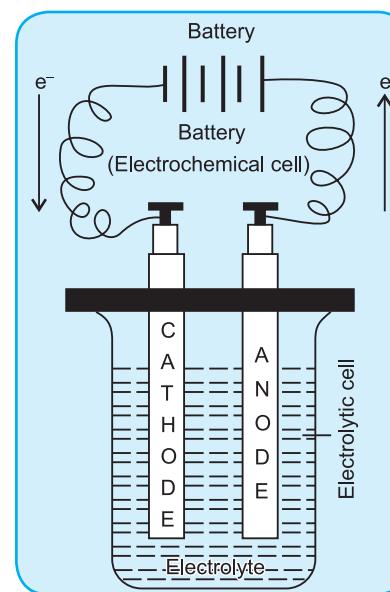


Fig. 6.1 : Electrolytic Cell

### Electrochemical Cell

It is a device used for carrying out the electrolysis experiment where the chemical energy is converted

into electrical energy. The different types of electrochemical cells are Daniel cell, simple voltaic cell etc.

### Ions

Ions are groups of atoms which carry a positive (+) or a negative charge (-) in it. There are two types of ions, cation and anion.

### Cation

The atom, which carries positive (+) charge on it is called cation.

### Anion

The atom which carries negative (-) charge on it is called anion.

**Fun Facts:** Number of articles are made using aluminium because of its light weight and its resistance to wear and tear against chemicals and weather. Its recycling became important due to the high demand. You can make 20 recycled aluminium cans from the energy needed to make one new can.

### Difference between cation and anions

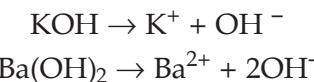
Cation	Anion
They are positively charged atoms.	They are negatively charged atoms.
During the electrolysis the cation moves toward cathode (-) since it is a negatively charged electrode.	During the electrolysis the anion moves towards anode (+) since it is a positively charged electrode.
Gain electrons from the cathode and are reduced to become a neutral atom. Example: $\text{Ca}^{2+}$ , $\text{Al}^{3+}$	Lose electrons to the anode and are oxidized to become a neutral atom. Example : $\text{Cl}^-$ , $\text{OH}^-$

### Ionic Compounds

Salts like sodium chloride, potassium chloride, magnesium chloride, calcium chloride, lead bromide, ammonium chloride and copper sulphate are examples of ionic compounds, which dissociate completely into anion and cation in solution or in molten state.



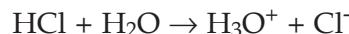
Soluble hydroxide of metals, i.e., strong bases, such as sodium hydroxide, potassium hydroxide and barium hydroxide are examples of compounds which dissociate completely into anion and cation.



### Strong Electrolytes

Strong acids dissociate completely into anion and cation in aqueous solutions like [HCl,  $\text{H}_2\text{SO}_4$  etc.]

### Hydrochloric Acid



### Sulphuric Acid



Hydrobromic acid and nitric acid are other examples for strong electrolytes.

### Weak Electrolytes

Electrolytes, which dissociate into ions in solution only to a small extent, are called weak electrolytes. Weak electrolytes are electrolytes that do not fully dissociate into ions in solution.

They exist as both **ions and molecules** in the solution due to their partial dissociation.

Weak electrolyte compounds are of the following type :

Weak inorganic acid, - e.g.,  $\text{H}_2\text{CO}_3$ ,  $\text{H}_2\text{SO}_3$  and  $\text{H}_3\text{PO}_4$



Organic acids, e.g.,  $\text{CH}_3\text{COOH}$  (acetic acid) and  $\text{HCOOH}$  (formic acid)



Weak bases, e.g.,  $\text{NH}_3$  and its derivatives



### Non-Electrolytes

Substances, which do not dissociate into ions in the molten state or in solution, are called non-electrolytes. They remain as molecules in the liquid state or solution and do not conduct electricity.

Examples of non-electrolytes are covalent compounds like pure water, pure acids, ethanol, benzene, carbon tetrachloride, sugar and paraffin wax. Pure water contains  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$  but to such a small extent ( $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 10^{-7} \text{ mol L}^{-1}$ ) that it does not conduct electricity.

## Chemical Changes that takes Place when an Electric Current Passes through an Electrolyte

### Experiment-1

Take some crystals of lead (II) bromide in a porcelain crucible, melt it, pass electric current through the molten salt, and observe it. You will find a silvery-grey material collects around the cathode and red-brown vapours evolve at the anode.

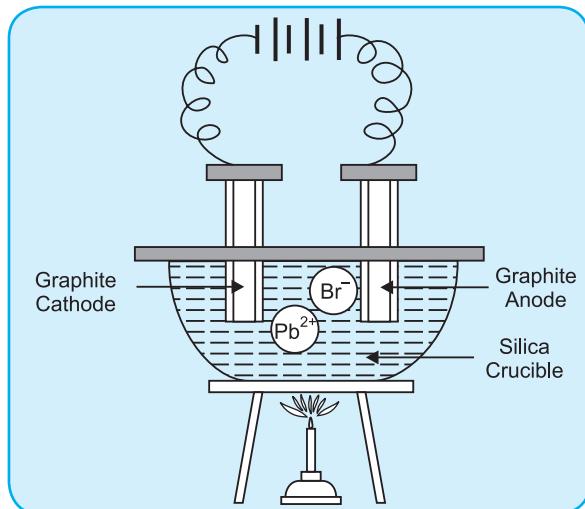


Fig. 6.2 : Electrolysis of Molten Lead Bromide

### Inference

Lead (II) bromide decomposes into lead, the silver-grey material collected at the cathode and bromine as the red-brown vapours evolved at the anode.

### Words to Know

**Anode :** The electrode in an electrolytic cell through which electrons move from the electrolyte to the battery.

**Cathode :** The electrode in an electrolytic cell through which electrons move from the battery to the electrolyte.

**Electrolyte :** Any substance that, when dissolved in water, conducts an electric current.

**Electrolytic cell :** A system in which electrical energy is used to bring about chemical changes.

**Electroplating :** A process that uses an electrolytic cell to deposit a thin layer of metal on some kind of surface.

**Ion :** Any particle, such as an atom or molecule, that carries an electric charge.

### Experiment-2

An aqueous solution of NaCl mixed with litmus solution is placed in a U – tube. When an electric current is passed through the electrolytic mixture, the

solution present in the cathode turns blue in colour and that in the anode turns white.

### Inference

A base sodium hydroxide has been produced at the cathode that turns the litmus blue and bleaching agent chlorine is released at the anode.

### Experiment-3

When an electric current is passed through acidified water or acidulated water, it decomposes into hydrogen and oxygen.

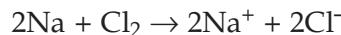
## 6.2 OXIDATION AND REDUCTION IN TERMS OF ELECTRON TRANSFER

Chemical reactions involving simultaneous oxidation and reduction are called 'redox reactions'. The oxidation and reduction in terms of electron transfer is based on the transfer of electrons from one atom to another atom. In terms of electron transfer, removal of one or more electrons from an atom is called oxidation whereas, addition of one or more electrons to an atom is called reduction. Therefore, a substance, which gives one or more electrons to the other, is called a reducing agent. And a substance which accepts one or more electrons from others is called an oxidizing agent.

### Formation of NaCl From Na and Cl<sub>2</sub>

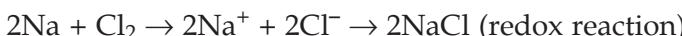
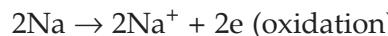
In the formation of NaCl from Na and Cl<sub>2</sub>, Na atom loses one electron to form a Na<sup>+</sup> ion and this process is called an oxidation process and the Cl atom gains that electron to form a Cl<sup>-</sup> ion and it is known as a reduction process. Therefore the formation of NaCl is a redox reaction. As Cl<sub>2</sub> molecule requires two electrons to form two Cl<sup>-</sup> ions, two Na atoms are ionized to give the electrons.

The reaction is represented as follows



Here during the oxidation reaction 2Na atoms lose electrons and become 2Na<sup>+</sup>.

And during the reduction reaction Cl<sub>2</sub> atoms gain electrons and become 2Cl<sup>-</sup>.



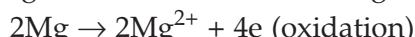
### Formation of MgO From Mg and O<sub>2</sub>

In this reaction, a Mg atom loses two electrons to form an Mg<sup>2+</sup> ion and an O atom gains them to form an O<sup>2-</sup> ion. Thus, the formation of Mg<sup>2+</sup> from Mg is an oxidation reaction and that of O<sup>2-</sup> from O is a reduction reaction.

Two Mg atoms are ionized to give four electrons to change one O<sub>2</sub> molecule into two O<sup>2-</sup> ions.



2Mg undergoes oxidation reaction to give 2Mg<sup>2+</sup>  
O<sub>2</sub> undergoes reduction reaction to give 2O<sup>2-</sup>

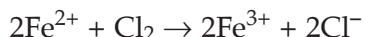


Mg acts as the reducing agent as it loses electrons and O<sub>2</sub> acts as the oxidizing agent as it gains electrons. At last the Mg<sup>2+</sup> and O<sup>2-</sup> ions combine to form MgO.

### Conversion of FeCl<sub>2</sub> To FeCl<sub>3</sub> By Cl<sub>2</sub>

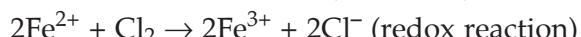
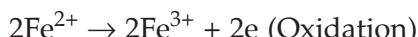


FeCl<sub>2</sub> is made up of one Fe<sup>2+</sup> ion and two Cl<sup>-</sup> ions. FeCl<sub>3</sub> is made up of one Fe<sup>3+</sup> ion and three Cl<sup>-</sup> ions. Hence, the reaction may be represented as:



Here 2Fe<sup>2+</sup> undergoes oxidation reaction and gives 2Fe<sup>3+</sup>

And Cl<sub>2</sub> undergoes reduction reaction and gives 2Cl<sup>-</sup>



Here, two Fe<sup>2+</sup> ions are oxidized to two Fe<sup>3+</sup> ions, and two Cl atoms (in Cl<sub>2</sub>) are reduced to form two Cl<sup>-</sup> ions. The Cl<sub>2</sub> molecule is the oxidizing agent and the Fe<sup>2+</sup> ion (or FeCl<sub>2</sub>) is the reducing agent.

### ELECTROLYTIC DISSOCIATION OF SODIUM CHLORIDE

In solid sodium chloride crystal, an electrostatic force of attraction holds Na<sup>+</sup> ion and Cl<sup>-</sup> ion together. It does not conduct electricity due to the absence of mobility of free ion.

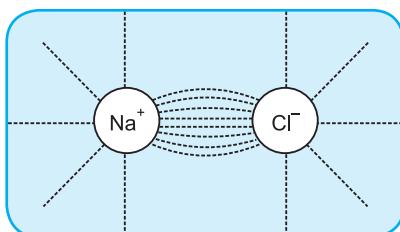


Fig. 6.3 : Electrostatic Force of Attraction between the Sodium Chloride Ions

In liquid and molten state the electrostatic force of attraction between sodium chloride ions that is Na<sup>+</sup>

and Cl<sup>-</sup> ions becomes weak. Thus, the bonds between the ions are broken and the ions are set free to move here and there. These free ions conduct electricity.

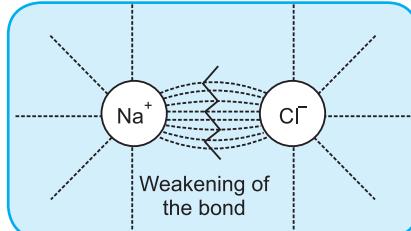


Fig. 6.4 : Weakening of the Bond between the Sodium Chloride Ions

### 6.3 ELECTROCHEMICAL SERIES OF METALS

Generally, most metallic elements tend to donate or lose electrons from their valence shell to form a positive metallic ion. This nature of donating electrons from the valence shell of metallic elements differs from metal to metal.

Sodium is a metallic element which is converted into a sodium ion by losing an electron. For e.g., Na → Na<sup>+</sup>. Similarly copper is converted into Cu<sup>2+</sup>. Zn is converted to Zn<sup>2+</sup> more easily than Ag to Ag<sup>+</sup>.

#### Electrochemical Series of Cations

Based on the ease with which atoms of metals lose electrons to form positively charged ions, the metals are arranged in a series known as the electrochemical series. The cation of a less active metal is discharged at the cathode in preference to that of a more active metal.

Table 6.1 Discharge of Cations

Ease of discharge of cations	
K	
Na	
Ca	
Mg	Increasing
Al	ease of
Zn	discharge
Fe	
H	
Cu	
Ag	↓

- In the electrochemical series, the metal in the higher position displaces the metal in the lower position.

- The metals that are placed below hydrogen cannot displace hydrogen from acid, but the metals placed above hydrogen can do it.
- Elements lower in the series get discharged more easily at the cathode during electrolysis because their cations can easily gain electrons for the elements. The metal ion accepts the electrons to form back the metal atom.
- Hydrogen has some properties of a metal, e.g., formation of a positive ion. Thus, its position amongst metals confirms its identity with metals. Therefore, it has been placed along with metals in the series.

### Electrochemical Series of Anions

Anions are also arranged in the increasing order of being discharged at the anode by losing electrons. This order is often called the electronegative series as they are negatively charged ions. During the electrolysis, they tend to lose their electron at the anode and become an atom. Not all the anions have identical tendency to lose electrons.

**Anions have been arranged in the increasing order, depending on the tendency to lose electrons and get discharged at the anode.**

**Table 6.2 Discharge of anion**

Ease of discharge of anions	
$\text{F}^-$	
$\text{SO}_4^{2-}$	Increasing ease of discharge
$\text{NO}_3^-$	
$\text{OH}^-$	
$\text{Cl}^-$	
$\text{Br}^-$	
$\text{I}^-$	↓

### 6.4 DISCHARGE OF IONS FROM AQUEOUS SOLUTIONS

During the electrolysis, the  $\text{H}^+$  ions migrate towards the cathode and  $\text{OH}^-$  ions towards the anode. Here the discharge of  $\text{H}^+$  ion and metal ion take place at the cathode and the discharge of anion of the salt and the  $\text{OH}^-$  ion at the anode.

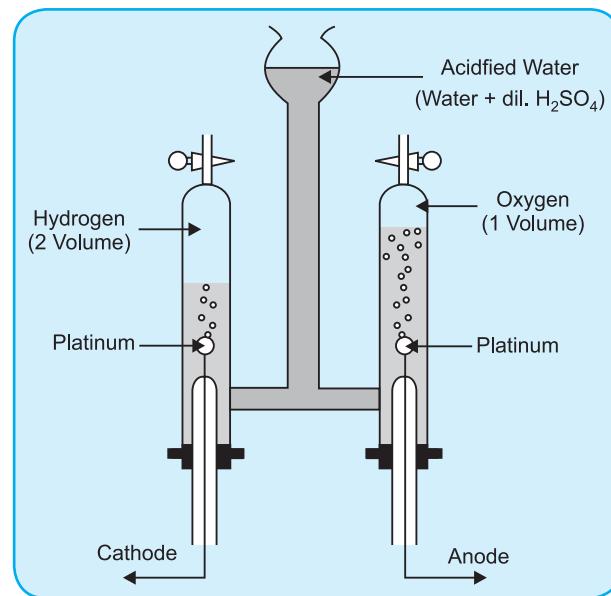
At cathode if the metal is more active than H, it will remain in ionic form in the solution, and  $\text{H}^+$  will be discharged. If the metal is less active than H, the metal ions will be discharged in preference to  $\text{H}^+$  ions, i.e.,  $\text{H}_3\text{O}^+$  ions and  $\text{H}^+$  or  $\text{H}_3\text{O}^+$  ions will remain in solution.

At anode the  $\text{Cl}^-$ ,  $\text{Br}^-$  and  $\text{I}^-$  is discharged in preference to  $\text{OH}^-$  and the products are  $\text{Cl}_2$ ,  $\text{Br}_2$  and  $\text{I}_2$  respectively.  $\text{OH}^-$  ions are discharged in preference to  $\text{SO}_4^{2-}$  and  $\text{NO}_3^-$  ions, and the product is  $\text{O}_2$ .

### Hofmann's Experiment

The Hofmann voltameter experiment is conducted for the electrolysis of acidified water.

The voltameter consists of two vertical graduated tubes namely A and B; its upper ends are fitted with stopcocks. These tubes are connected to each other by a horizontal tube, which in turn is connected to a long vertical funnel. The platinum electrodes are introduced into the vertical funnel and connected respectively to the negative and positive terminals of a battery. The acidified water is poured through the funnel into the voltameter and the current is switched on.



**Fig. 6.5 : Electrolysis of Acidified Water Using Platinum Electrode**

### Observation

We can find two volumes and one volume of gas collected in both the tubes A and B respectively. The gas collected at cathode burns with the pop sound so it is  $\text{H}_2$  gas.

The gas collected at the anode burns with the glowing taper so it is  $\text{O}_2$  gas.

### Aqueous Solution Of Copper (II) Sulphate Using Copper Electrodes

For an electrolysis experiment an aqueous solution of copper (II) sulphate is acidified with little amount of dilute sulphuric acid to increase the conductivity of the electrolyte.

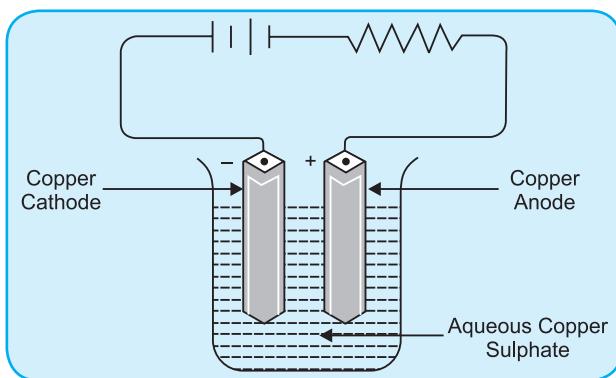
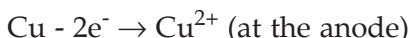
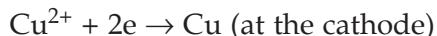


Fig. 6.6 : Electrolysis of Copper Sulphate Solution

Here two copper electrodes are used,



Thus, copper slowly dissolves out from the anode. The anode is made up of same metal as the metal ion of the electrolyte; the metal dissolves out of the anode and forms ions by giving up electrons to it. As a result of this reaction a brown color metal is deposited at the cathode electrode whereas the anode electrode becomes thin. During the electrolysis, the total number of  $\text{Cu}^{2+}$  ions remains the same because when a  $\text{Cu}^{2+}$  ion is discharged at the cathode, a  $\text{Cu}^{2+}$  ion is added to the solution by the anode.



## 6.5 APPLICATIONS OF ELECTROLYSIS

Electrolysis has many industrial applications as follows :

- In the production of metals such as sodium and potassium.
- In the production of gases such as hydrogen, oxygen, chlorine etc.,

## Electroplating

An important application of electrolysis is electroplating. The process of coating one metal by another metal by electrolysis is called electroplating. In this process, a superior metal like chromium, gold, silver, or nickel is coated (deposited) over a metal like copper or iron using electrolysis.

- Electroplating helps to serve the following purpose.

## Decoration

Metals used for making jewellery, e.g., gold and silver, are very expensive. In order to reduce the cost, cheaper metals are used to make jewels, and then coated with gold or silver.

## Protection

Metals that corrode easily are given a protective coating with a metal that has a lower tendency to corrode. Nickel and chromium plating are widely used in the automobile industry for this purpose.

- Electroplating is used in the manufacture of printed circuit boards (small electrical circuits, with electronic components generally printed on a plastic board), which are used in many appliances like radios, TV, computers, etc.

## Purification of metals

The purification of metals by the process of electrolysis is called electrolytic refining. Electrolysis is used in the preparation of very pure metals from impure metals.

- Electrodialysis is a method used for the conversion of sea water into drinking water.

## Products of Electrolysis

Table 6.3 Product of Electrolysis

Electrolyte	Products	
	Cathode	Anode
Aq. Sol. of dil $\text{H}_2\text{SO}_4$	Hydrogen	Oxygen
Aq. sol. of dil $\text{NaOH}$	Sodium	Oxygen
Aq. sol. of dil $\text{NaCl}$	Sodium	Chlorine
Aq. Sol. of con $\text{NaCl}$	Hydrogen	Chlorine
Aq. Sol. of $\text{CuSO}_4$	Copper	Oxygen
Molten sodium chloride	Sodium	Chlorine
Molten lead bromide	Lead	Bromine

## 6.6 ELECTROPLATING

Electroplating is a method by which a metal object can be coated with another metal by electrolysis.

Various objects are electroplated to protect them from corrosion and give them an attractive appearance. Some examples are :

Table 6.4 Applications of Electroplating

Electroplating Metal	Applications
Chromium	Water taps, car and bicycle parts
Nickel	Motor cycle parts, bathroom fittings
Silver	Silver trophies, plates, knives
Tin	Tin cans
Gold	Ornaments

### Electroplating With Nickel

Steel is electroplated with nickel to prevent rusting. An aqueous solution of the double salt nickel ammonium sulphate,  $(\text{NH}_4)_2\text{SO}_4 \cdot \text{NiSO}_4 \cdot 6\text{H}_2\text{O}$ , is used as the electrolyte for this purpose. In addition, its bluish green colour crystals are highly soluble in water.

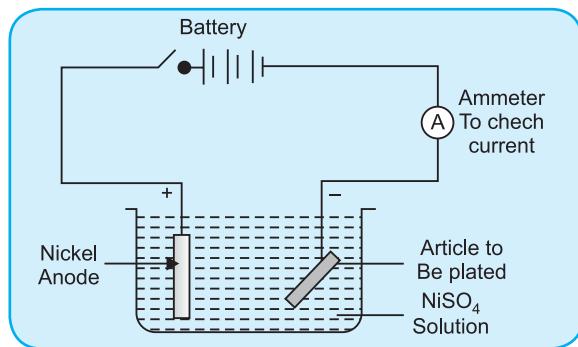
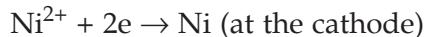


Fig. 6.7 : Electroplating of an Article with Nickel

The anode is made by a nickel bar and the cathode is made with the article to be plated (steel). During the electrolysis process, the  $\text{Ni}^{2+}$  ions are discharged and deposited slowly over the cathode on the steel. At the same time, nickel dissolves out of the anode and replenishes the  $\text{Ni}^{2+}$  ions in the electrolyte.



The  $\text{SO}_4^{2-}$  and  $\text{OH}^-$  ions move towards the anode, but neither is discharged.

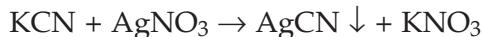
For the uniform deposition of electroplating a steady flow of current is passed.

### Electroplating With Silver

Silver plating is widely used for electroplating of cutlery and some fancy articles made of nickel silver or German silver, an alloy of Cu, Zn, and Ni.

$\text{AgNO}_3$  can be used as an electrolyte, but a coordination compound of the formula  $\text{K}[\text{Ag}(\text{CN})_2]$  is preferred to  $\text{AgNO}_3$ . The compound  $\text{K}[\text{Ag}(\text{CN})_2]$  is commonly known as potassium argentocyanide [potassium dicyanoargentate (I)].

An aqueous solution of  $\text{AgNO}_3$  is treated with one of  $\text{KCN}$ , and a white precipitate of  $\text{AgCN}$  is formed.  $\text{KCN}$  solution is again added till  $\text{AgCN}$  dissolves to form the argentocyanide.



A white precipitate of silver cyanide is formed

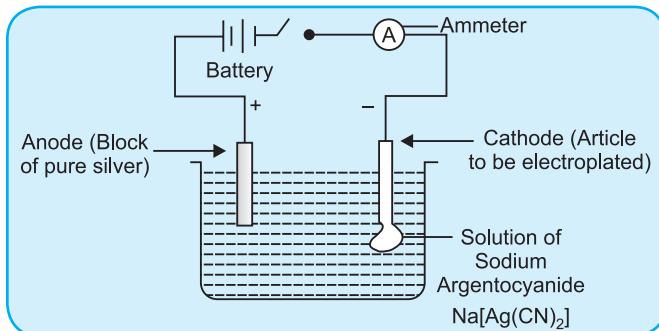
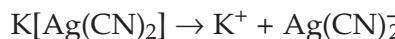
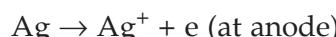


Fig. 6.8 : Electroplating of a Brass Spoon with Silver

The silver is made as anode and the article to be plated is made the cathode. The electrolyte dissociates as follows.



Upon electrolysis, Ag is deposited at the cathode and the same amount of the metal dissolves out of the anode.



Passage of steady current deposits the metal coating uniformly on the article during electroplating.

## 6.7 ELECTROREFINING

### Purification of Copper

This is another important industrial use of electrolysis. An interesting application is the purification of copper metal. When the copper metal is separated from its ore, it is 99% pure. The other impurities are iron and zinc. Impure copper is made as the anode and pure copper is the cathode. The electrolyte is  $\text{CuSO}_4$  solution.

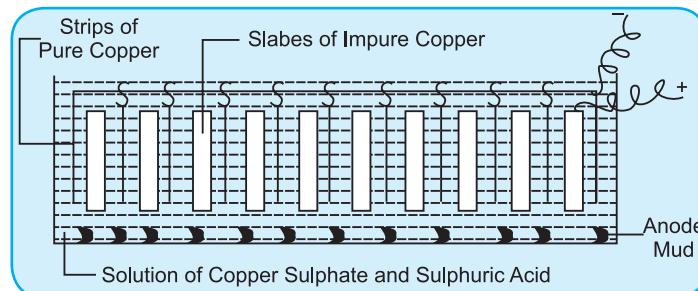
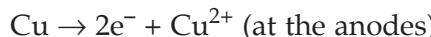
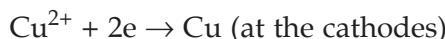


Fig. 6.9 : Refining of Copper

When electrolysis is carried out, the copper ions gain electrons and they are deposited on the pure

copper bar. The other impurities in the crude copper settle down at the bottom of the solution. It is collected below the anode as anode mud. At the cathode,  $\text{Cu}^{2+}$  ions react to form copper.

### Acids, Bases and Salts as Electrolytes

Acids like  $\text{HCl}$ ,  $\text{HNO}_3$  and  $\text{H}_2\text{SO}_4$  give  $\text{H}_3\text{O}^+$  and the corresponding anions in aqueous solution. Bases like  $\text{NaOH}$  and  $\text{KOH}$ , which do not have mobile electrons can also conduct electricity in the molten state or in an aqueous solution. Under these conditions, they form mobile ions, which carry charge across them.

Salts like  $\text{NaCl}$  and  $\text{CuSO}_4$  dissociate into cations and anions in the molten state as well as in solution.

Generally acid, base and salts are categorized as strong electrolyte and weak electrolyte based on the degree of dissociation. i.e.,

$$\text{Degree of dissociation} \propto =$$

$$\frac{\text{number of molecules dissociated}}{\text{total number of molecules}} \times 100$$

This is the order of tendency of metals and hydrogen to form cations. So, the cation of a less reactive metal is discharged at the cathode in preference to that of a more active metal.

$\text{Mg}$  being lower than  $\text{Na}$  in the activity series,  $\text{Mg}^{2+}$  ions will be discharged at the cathode in preference to  $\text{Na}^+$  ions from a mixture of the two. Similarly,  $\text{Fe}^{2+}$  ion will be discharged in preference to  $\text{Mg}^{2+}$  ions, and  $\text{Cu}^{2+}$  ions in preference to  $\text{Fe}^{2+}$  or  $\text{H}^+$  ions at the cathode.

### INTEXT QUESTIONS

1. Define electrolysis.
2. Define cation and anion.
3. What is the difference between weak electrolyte and non-electrolyte ?
4. Name the salt which is a weak electrolyte and a base which is a weak electrolyte.
5. Explain Arrhenius theory of electrolysis.
6. Explain why sodium chloride crystal does not conduct electricity but the molten or liquid state of sodium chloride conducts electricity.
7. Explain how hydrochloric acid is a good conductor of electricity.
8. Why cannot we store silver nitrate solution in copper vessel ?
9. How would you change a metal like Cu into its ions ?
10. How would you change  $\text{Cu}^{2+}$  ion to Cu ?
11. An electrolytic cell is set up using two platinum electrodes and an aqueous solution of copper (II) sulphate.
  - (a) Draw a labeled diagram of the electrolytic cell.
  - (b) Name the ions present in the cell.
  - (c) Name the ions migrating towards the anode.
  - (d) Name the ions migrating towards the cathode.
  - (e) Name the ions which will not be discharged at electrodes during electrolysis.
  - (f) Write the reaction at the cathode and anode.
  - (g) Name the spectator ion in the solution.



### POINTS TO REMEMBER

- ◆ The passage of electric current through matter is called electrical conduction.
- ◆ Metallic or electronic conductors conduct electric current without undergoing a chemical change. All metals, alloys and non-metals like graphite can conduct electric current.
- ◆ The decomposition of a compound when electric current is passed through it in the molten state or in solution is known as electrolysis.
- ◆ Substances which allow the passage of electric current through them in the molten state or in solution due to the presence of ions are called electrolytes.
- ◆ Substances that do not dissociate into ions in the molten state or in aqueous solutions are called non-electrolytes.
- ◆ Cation is a positive ion and anion is a negative ion.
- ◆ Anode is a positive electrode and cathode is negative electrode.
- ◆ Ionic compounds act as electrolytes in the molten state as well as in aqueous solutions.
- ◆ Some polar covalent compounds like acids also act as electrolytes, but only in aqueous solutions.

- ◆ Electrolytes which get completely dissociated into ions in aqueous solution are known as strong electrolytes.
- ◆ Electrolytes which get dissociated into ions in aqueous solutions only to a small extent are known as weak electrolytes.
- ◆ The substances which neither in solution nor in the molten state allows an electric current to pass through it are non-electrolytes. It contains only molecules. e.g., pure water, pure acid, benzene, and ethanol.

- ◆ An element having a smaller tendency to form ions is deposited at an electrode in preference to one having a greater tendency to form ions.
- ◆ The cation of a less active metal is discharged at the cathode in preference to that of a more active metal.
- ◆ Electrolysis of acidified water produces hydrogen and oxygen.
- ◆ Coating of a metal on another metal by electrolysis is known as electroplating.
- ◆ The process of purification of metals by employing electrolysis is called electrorefining.

### EXERCISES

#### A. FILL IN THE BLANKS

1. Sodium hydroxide is.....
2. Salts do not conduct electricity in the .....state
3. Pure water consist almost entirely of ..... [ICSE 2002]
4. Pure water.....conduct electricity [ICSE 2002]
5. During the electrolysis of acidified water using platinum electrodes, hydrogen is liberated at the.....and oxygen at the ..... [ICSE 2002]
6. Electrolysis is the passage of ....through a liquid or a solution accompanied by a .... change.
7. Complete the following table, which refers to two practical applications of electrolysis. [ICSE 2003]

	Anode	Electrolyte	Cathode
Silver plating of a spoon		Solution of potassium argentocyanide	
Purification of copper			

8. As we descend the electrochemical series containing cations, the tendency of the cations to get.... at the cathode increases. [ICSE 2005]
9. The .....the concentration of an ion in a solution, the greater is the probability of its being discharged at the appropriate electrode. [ICSE 2005]

#### B. MULTIPLE CHOICE TYPE QUESTIONS

Choose the Correct Answer from the options given below :

1. Which of the following substances is an electrolyte ?
 

(a) Mercury	(b) Copper
(c) Sodium sulphate	(d) Aluminium
2. Which of the following species will be deposited at the cathode on the electrolysis of an aqueous solution of potassium bromide ?
 

(a) K	(b) H <sub>2</sub>
(c) Br <sub>2</sub>	(d) O <sub>2</sub>
3. If you want to electrolyse concentrated HCl, which of the following will you choose for making the anode ?
 

(a) Graphite	(b) Aluminium
(c) Iron	(d) Copper
4. For electroplating a brass jug with copper, the anode must be made of :
 

(a) Zinc	(b) Copper
(c) Nickel	(d) Platinum
5. During the electrolysis of acidified water which of the following takes place
 

(a) Oxygen is released at cathode.
(b) Oxygen is released at anode.
(c) Hydrogen is released at anode.
(d) Sulphur dioxide is released at anode.

#### C. GIVE APPROPRIATE SCIENTIFIC REASONS FOR THE FOLLOWING STATEMENTS

- (i) Carbon tetrachloride does not conduct electricity.
- (ii) During electrolysis of molten lead bromide graphite anode is preferred to other electrodes.



- (iii) The electrical conductivity of acetic acid is less in comparison to the electrical conductivity of dilute sulphuric acid at a given concentration.
- (iv) Electrolysis of molten lead bromide is considered to be a redox reaction. [2015]

#### D. MATCH THE COLUMNS

Column I	Column II
1. Weak electrolyte	Bathroom fittings
2. Tin	Ornaments
3. Silver	Water caps
4. Gold	Knives
5. Chromium	Tin cans
6. Nickel	$\text{CH}_3\text{COOH}$

#### E. STATE ONE RELEVANT OBSERVATION

- (i) At the Anode when aqueous copper sulphate solution is electrolysed using copper electrodes. [2015]

#### F. VERY SHORT QUESTIONS

- Give one example each of a substance, which contains (i) ions only, (ii) molecules only, and (iii) ions as well as molecules. [ICSE 1996]
- Classify the following substances into strong electrolytes, weak electrolytes and non electrolytes. (a) acetic acid (b) ammonium sulphate (c) ammonium hydroxide (d) carbon tetrachloride (e) sodium acetate (f) dilute hydrochloric acid (g) dilute sulphuric acid [ICSE 2002]
- What should be the physical state of lead bromide if it is to conduct electricity? [ICSE 1997]
- In the formation of magnesium chloride (by direct combination between magnesium and chlorine), name the substance that is oxidised and the substance that is reduced. [ICSE 2007]
- Name the products at the cathode and the anode when molten lead bromide is electrolysed. [ICSE 1997, 1998]
- If fused metallic chloride is electrolysed, at which electrode will the metal be obtained? [ICSE 1999]
- $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}$   
At which electrode (anode or cathode) would such a reaction take place? Is this an example of oxidation or reduction? [ICSE 2008]

#### G. SHORT QUESTIONS

- Define the term electrolyte. [ICSE 1996]
- What is a non-electrolyte? [ICSE 2003]
- Why carbon tetrachloride liquid does not conduct electricity?
- During the electroplating of silver, silver nitrate is not used as electrolyte. Explain?
- Define a strong electrolyte with an example.
- Define a weak electrolyte with an example.
- Why does the addition of sulphuric acid to water produce a conducting solution? [ICSE 1995] [ICSE 2002]
- Hydrogen chloride and ethanol are covalent compounds soluble in water. The aqueous solution of which of the two will conduct electric current? Write down the relevant chemical equation.
- What do you mean by electrolysis? [ICSE 1992] [ICSE 1995]
- Acidified water is electrolysed using platinum electrodes.
  - What is the volume ratio of hydrogen to oxygen?
  - Give the electrode reactions. [ICSE 2002]
- An aqueous solution of nickel sulphate contains  $\text{Ni}^{2+}$  and  $\text{SO}_4^{2-}$  ions.
  - Which ion moves toward the cathode?
  - What is the product at the anode?
- When an aqueous solution of copper (II) sulphate is electrolysed with copper electrodes, the blue copper ions are discharged at the cathode and are deposited as copper, yet the blue colour of the solution does not fade. Explain. [ICSE 1992]
- Give the electrode reactions that occur when an aqueous solution of sodium chloride is electrolysed.
- Name the three industrial applications of electrolysis. [ICSE 1992]
- During the electrolysis of copper(II) sulphate solution using platinum as cathode and carbon as anode:
  - What do you observe at the cathode and at the anode?
  - What change is noticed in the electrolyte?
  - Write the reactions at the cathode and at the anode.
- Answer the following questions related to the silver plating of a copper article.
  - Which electrode is formed by the article to be plated?

- (b) Which ions must be present in the electrolyte?
- (c) Of what material should the anode be made?
- (d) Give the cathode reaction.

[ICSE 1994, 1995]

17. A metal article is to be electroplated with silver. The electrolyte selected is sodium argentocyanide.
- (a) What kind of salt is sodium argentocyanide?
  - (b) Why is it preferred to silver nitrate as an electrolyte?
  - (c) State one condition to ensure that the deposit is smooth, firm and long lasting.
  - (d) Write the reaction taking place at the cathode.
  - (e) Write the reaction taking place at the anode.

[ICSE 2009]

18. A solution contains magnesium ions ( $Mg^{2+}$ ), iron (II) ions ( $Fe^{2+}$ ) and copper ions ( $Cu^{2+}$ ). On passing an electric current through this solution, which ions will be the first to be discharged at the cathode? Write the equation for the cathode reaction.

[ICSE 2008]

## H. LONG QUESTIONS

1. What are the practical applications of the electrolysis of an aqueous solution of copper (II) sulphate? Describe one such application.

[ICSE 1996]

2. Explain how electrolysis is an example of a redox reaction.

[ICSE 2004]

3. Mr. Ramu wants to electroplate his key chain with nickel to prevent rusting. For this electroplating:

- (a) Name the electrolyte.
- (b) Name the cathode.
- (c) Name the anode.
- (d) Give the reaction at the cathode.
- (e) Give the reaction at the anode.

[ICSE 2010]

4. (a) Draw a neat and well labelled diagram for the silver plating on an iron spoon.
- (b) Complete the following table related to electrolysis.

Sl. no.	Name of Electrolyte	Name of cathode	Name of anode	Product at cathode	Product at anode
1	$CuSO_4$ (aq)	Copper	Copper		
2	$PbBr_2$ (molten)	Platinum	Platinum		

- (c) Classify the following as oxidation and reduction reaction, also complete the reaction.
- (i)  $Cu \rightarrow Cu^{2+}$
  - (ii)  $Fe^{3+} \rightarrow Fe^{2+}$
  - (iii)  $Cl^- \rightarrow Cl$

[ICSE 2012]

## RECENT BOARD QUESTIONS

### Question 1.

- (a) Give appropriate scientific reasons for the following statements : [2015]

- (i) Zinc oxide can be reduced to zinc by using carbon monoxide, but aluminium oxide cannot be reduced by a reducing agent.
- (ii) Carbon tetrachloride does not conduct electricity.
- (iii) During electrolysis of molten lead bromide graphite anode is preferred to other electrodes.
- (iv) The electrical conductivity of acetic acid is less in comparison to the electrical conductivity of dilute sulphuric acid at a given concentration.
- (v) Electrolysis of molten lead bromide is considered to be a redox reaction.
- (vi) At the Anode when aqueous copper sulphate solution is electrolysed using copper electrodes.

- (b) (i) For each of the substance listed below, describe the role played in the extraction of aluminium. [2015]

- (1) Cryolite
- (2) Sodium hydroxide
- (3) Graphite

- (c) Select the correct answer from the list given in brackets : [2012]

- (i) An aqueous electrolyte consists of the ions mentioned in the list, the ion which could be discharged most readily during electrolysis.  
[ $Fe^{2+}, Cu^{2+}, Pb^{2+}, H^+$ ].
- (ii) The metallic electrode which does not take part in an electrolytic reaction. [Cu, Ag, Pt, Ni].
- (iii) The ion which is discharged at the Cathode during the electrolysis of copper sulphate solutions using copper electrodes as anode and cathode. [ $Cu^{2+}, OH^-, SO_4^{2-}, H^+$ ].
- (iv) When dilute sodium chloride is electrolysed using graphite electrodes, the cation is discharged at the cathode most readily.



- (v) During silver plating of an article using potassium argentocyanide as an electrolyte, the anode material should be [Cu, Ag, Pt, Fe].

**(d) Explain why :** [2011-15,16]

- In the electrolysis of alumina using the Hall Heroult's Process the electrolyte is covered with powdered coke.
- Iron sheets are coated with zinc during galvanization.
- In the electroplating of an article with silver, the electrolyte sodium argentocyanide solution is preferred over silver nitrate solution.
- The electrolysis of acidulated water is considered to be an example of catalysis.

(ii) Copy and complete the following table :

	Anode	Electrolyte
Purification of copper		

(b) Write the equation taking place at the anode.

**Question 2.**

(a) During the electrolysis of copper (II) sulphate

solution using platinum as cathode and carbon as anode : [2011]

- What do you observe at the cathode and at the anode.
- What change is noticed in the electrolyte.
- Write the reactions at the cathode and at the anode.

**Question 3.**

(a) Write equations for the reactions taking place at the two electrodes (mentioning clearly the name of the electrode) during the electrolysis of : [2016]

- Acidified copper sulphate solution with copper electrodes.
- Molten lead bromide with inert electrodes.

(b) Answer in one word : [2016]

- Name the product formed at the anode during the electrolysis of acidified water using platinum electrodes.

- Name the metallic ions that should be present in the electrolyte when an article made of copper is to be electroplated with silver.

- Name the compound added to pure alumina to lower the fusion temperature during the electrolytic reduction of alumina.



# 7

# METALLURGY

## LEARNING OUTCOMES

- 7.0. Introduction
- 7.1. The Position of Metals and Non-metals in the Periodic Table
- 7.2. Comparative Study of the Properties of Metals and Non-metals
- 7.3. Corrosion
- 7.4. Extraction of Metals Based on the Activity Series
- 7.5. Some Common Terms Used in Extraction

- 7.6. Steps Involved in Extraction
- 7.7. Calcination and Roasting
- 7.8. Aluminium
- 7.9. Properties of aluminium
- 7.10. Uses of Aluminium
- 7.11. Forms and Uses of Iron
- 7.12. Uses of Zinc
- 7.13. Alloys and Amalgam
- 7.14. A List of Some Common Alloys along with their Composition and Uses

## 7.0 INTRODUCTION

We always arrange our books in our racks according to the subjects, size, etc. Similarly, there are about 118 elements known to us today, which are arranged in the periodic table according to their size, atomic number, mass number, physical and chemical properties, etc.

These elements are widely distributed on the earth's crust, both in the free state as well as in a combined form. Out of these known elements, more than 80 are metals and 22 are non-metals.

All the metallic elements are placed at the left, in the middle and at the bottom of the periodic table; whereas, non-metals are placed on the right-hand side of the periodic table.

All metals (except mercury, which is a liquid) in the periodic table are solids at room temperature;

whereas, out of 22 non-metals, 11 are gases, 1 is liquid (bromine) and the remaining 10 non-metals are solids at room temperature.

### Metals

➤ **Metals are elements that are lustrous crystalline solids of high strength and high melting and boiling points; they are good conductors of heat and electricity; they are malleable, ductile and electropositive in nature.**

Metals generally used in our daily life are iron, copper, tin, lead, silver, nickel and mercury. Metals are found to occur in nature either in the native or free state or combined state.

Metals such as silver, gold, platinum, etc., are less reactive and occur in the native state; whereas, metals such as copper, aluminium, sodium etc., are highly active and occur in the combined state with halogens,

sulphur, oxygen, etc. For e.g., copper pyrites ( $\text{CuFeS}_2$ ), bauxite ( $\text{Al}_2\text{O}_3$ ), common salt ( $\text{NaCl}$ ), etc.

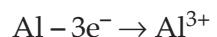
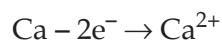
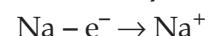
### Biological Application

- K and Na are present in body fluids.
- Mg and S are found in proteins.
- Ca and P are present in teeth and bones.

Each major mineral is present in 0.1% – 2% by mass. At least 100 mg of each mineral is needed for human consumption in daily life.

Most of the trace metals such as As, B, Cr, Co, Cu, Fe, Mn, Ni, Zn etc. are present in less than 0.1% by mass. A small quantity (15 mg or <) of each element is needed in one's daily diet.

*Chemically, metals are electropositive in nature because they have the tendency to lose electrons easily and form positively charged ions (cations). For example, sodium, magnesium and aluminium are metals because they form positive ions by the loss of 1, 2 and electrons, respectively.*



### Non-Metals

- Non-metals are elements that are not lustrous, are bad conductors of heat and electricity, brittle if solid, and electronegative in nature.

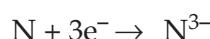
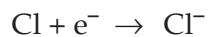
Hydrogen, carbon, oxygen, sulphur and phosphorus are some non-metals, which are quite abundant on the earth. Hydrogen, oxygen, carbon, sulphur, silicon, phosphorous and inert gases are some of the common non-metals.

### Biological Application

Building-block elements

O, C, H, and N are four elements that compose almost 96% of the mass of the human body.

*Chemically, non-metals are electronegative in nature because they have the tendency to gain electrons and form negatively charged ions (anions). For example, chlorine, oxygen and nitrogen are non-metals because they form negative ions by gaining 1, 2 and 3 electrons, respectively.*



## 7.1 THE POSITION OF METALS AND NON-METALS IN THE PERIODIC TABLE

### Metals

- Metals occupy a large area of the periodic table and they are present on the left hand side of the staircase demarcation in the periodic table.
- Alkali metals like lithium, sodium, potassium, rubidium, caesium and francium belongs to group IA which has one valence electron in the outer shell of their atom.
- Alkaline earth metals like beryllium, magnesium, calcium, strontium, barium and radium belong to group IIA, which has two valence electrons in the outermost orbit of the atom.
- Transition metals like iron, zinc, copper and mercury are placed in group 3-12 of the periodic table. They have two outermost incomplete shells.
- Inner transition metals belong to group IIIB and in the period six. They are called lanthanides. The seventh period is called actinides.
- Metals such as Al, Sn and Pb are present in the group of the lower part 13 – 17.
- All these metals have similar properties based upon the groups they belong.

**Table 7.1 Position of metals in the periodic table**

Group	Metals	Position
IA	Alkali metals	First column on the left side of the periodic table
IIA	Alkaline earth metals	Second column on the left side of the periodic table
3 - 12	Transition metals	Middle of the periodic table
13 – 17	Lanthanides actinides	Bottom of the periodic table

### Non-Metals

Non-metals are placed on the extreme right side of the periodic table, except the element hydrogen, which is placed on the left side of the periodic table. Non-metals have usually 4 to 8 electrons in the outermost shells of their atoms.

For example, carbon, nitrogen, fluorine and neon are non-metals whose atomic numbers are 6, 7, 8, 9 and 10 respectively. Hydrogen is a non-metal having one electron in the outermost shell of its atom and

helium is another non-metal having two electrons in the outermost shell of its atom. Their electronic configurations are as follows :

**Table 7.2 Electronic configuration of non-metals**

Element	K	L
Carbon	2	4
Nitrogen	2	5
Oxygen	2	6
Fluorine	2	7
Neon	2	8

#### NOTE

In the periodic table, element 110 is Ds, 111 is Rg, and 112 is Cn. Element 114 does not exist.

In sixteenth century people know that 'Silver salts' darken when exposed to light because of the formation of tiny particles of metallic silver. When silver chloride or bromide precipitates are kept exposed to sunlight, they will darken after few hours.

Here silver ions are converted into silver atoms. This principle was later used as the basis of black-and-white and coloured film photography.

#### Metal Detectors

Metal detectors can find metallic objects and coins buried under soil and are used for security reasons to detect weapons such as knives or guns. A metal detector has an oscillator producing an alternating current that passes through a coil producing an alternating magnetic field. When a metal object, which is electrically conductive, is close to this coil, eddy current get induced in the metal thereby producing an alternating magnetic field of its own. Another coil is placed to measure this change in the magnetic field produced to which the metallic object can be detected.

While aluminum is the most common metal found in the earth's crust, the most common metal found on earth is iron, mostly because it makes up such a large part of the earth's core followed by magnesium.

# PERIODIC TABLE OF THE ELEMENTS

1 1.008 <b>H</b> HYDROGEN	2 4.0026 <b>He</b> HELIUM
3 6.94 <b>Li</b> LITHIUM	4 9.0122 <b>Be</b> BERYLLIUM
11 22.990 <b>Na</b> SODIUM	12 24.305 <b>Mg</b> MAGNESIUM
19 39.098 <b>K</b> POTASSIUM	20 40.078 <b>Ca</b> CALCIUM
37 85.468 <b>Rb</b> RUBIDIUM	38 87.62 <b>Sr</b> STRONTIUM
55 132.91 <b>Cs</b> CAESIUM	56 137.33 <b>Ba</b> BARIUM
87 (223) <b>Fr</b> FRANCIUM	88 (226) <b>Ra</b> RADIUM
57 138.91 <b>La</b> LANTHANUM	58 140.12 <b>Ce</b> CERIUM
89 (227) <b>Ac</b> ACTINIUM	90 232.04 <b>Th</b> THORIUM
6 12.011 <b>B</b> BORON	7 14.007 <b>C</b> CARBON
13 26.982 <b>Al</b> ALUMINUM	14 28.085 <b>Si</b> SILICON
21 44.956 <b>Sc</b> SCANDIUM	22 47.867 <b>Ti</b> TITANIUM
39 88.906 <b>Y</b> YTTRIUM	40 91.224 <b>Zr</b> ZIRCONIUM
41 92.906 <b>Nb</b> NIOBIUM	42 95.95 <b>Mo</b> MOLYBDENUM
43 (98) <b>Tc</b> TECHNETIUM	44 101.07 <b>Ru</b> RUTHENIUM
46 106.42 <b>Pd</b> PALLADIUM	47 107.87 <b>Ag</b> SILVER
49 114.82 <b>Ga</b> GALLIUM	50 118.71 <b>Cd</b> CADMIUM
51 121.76 <b>Ge</b> GERMANIUM	52 127.60 <b>In</b> INDIUM
53 126.90 <b>As</b> ARSENIC	54 131.29 <b>Se</b> SELENIUM
55 127.60 <b>Br</b> BROMINE	56 131.29 <b>Te</b> TELLURIUM
57 126.90 <b>Kr</b> KRYPTON	58 131.29 <b>I</b> IODINE
59 126.90 <b>Xe</b> XENON	60 131.29 <b>Po</b> POLONIUM
61 131.29 <b>At</b> ASTATINE	62 131.29 <b>Pb</b> LEAD
63 131.29 <b>Bi</b> BISMUTH	64 131.29 <b>Tl</b> MERCURY
65 131.29 <b>Dy</b> DYPROSIMUM	66 131.29 <b>Ho</b> ERBIUM
67 131.29 <b>Tm</b> THULIUM	68 131.29 <b>Er</b> YTERBIUM
69 131.29 <b>Yb</b> YTTERBIUM	70 131.29 <b>Lu</b> LUTETIUM
71 131.29 <b>Fm</b> FERMIUM	72 131.29 <b>Md</b> MENDELEVIIUM
73 131.29 <b>No</b> NOBELIUM	74 131.29 <b>Og</b> OGANESSON
75 131.29 <b>Fr</b> FRANCIUM	76 131.29 <b>Ts</b> TENNESSEINE
77 131.29 <b>Rg</b> RADIUM	78 131.29 <b>Lv</b> LIVERMORIUM
79 131.29 <b>Cn</b> COPERNICIUM	80 131.29 <b>Mc</b> MUSCOMUM
81 131.29 <b>Uut</b> UNUNTRIUM	82 131.29 <b>Fl</b> FLEROMIUM
83 131.29 <b>Th</b> THORIUM	84 (209) <b>Pa</b> PROTACTINIUM
85 (210) <b>Rn</b> RADON	86 (222) <b>At</b> ASTATINE
87 (223) <b>Fr</b> FRANCIUM	88 (226) <b>Ra</b> RADIUM
89 (227) <b>Ac</b> ACTINIUM	90 232.04 <b>Th</b> THORIUM
91 231.04 <b>Pa</b> PROTACTINIUM	92 238.03 <b>U</b> URANIUM
93 (237) <b>Np</b> NEPTUNIUM	94 (244) <b>Pu</b> PLUTONIUM
95 (243) <b>Cm</b> AMERICIUM	96 (243) <b>Am</b> CURIUM
97 (247) <b>Bk</b> BERKELEIUM	98 (247) <b>Cf</b> CALIFORNIUM
99 (252) <b>Fm</b> EINSTEINIUM	100 (257) <b>Md</b> FERMIUM
101 (258) <b>No</b> MENDELEVIIUM	102 (259) <b>Es</b> OXYGEN
103 (262) <b>Lr</b> LAWRENCIUM	104 (263) <b>Og</b> OGANESSON

\* Lanthanide series

\*\* Actinide series

## 7.2 COMPARATIVE STUDY OF THE PROPERTIES OF METALS AND NON-METALS

**Table 7.3 Physical Properties of Metals and Non-Metals**

Physical Properties	Metals	Non-metals
<b>1. Physical state</b>	<i>All metals are solids at room temperature. Exception: Mercury is a liquid at room temperature while gallium changes to the liquid state at slightly above the room temperature.</i>	Non-metals exist in all the three states. They are usually solids or gases. <i>Exception: Bromine is the only non-metal that exists in a liquid state.</i>
<b>2. Hardness</b>	Metals are generally hard and strong. <i>Exception: Sodium and potassium are soft metals and can be cut with a knife.</i>	Non-metals are generally soft. <i>Exception: Carbon in the form of diamond is the hardest object known.</i>
<b>3. Metallic lustre</b>	Metals possess a bright, shiny, metallic luster. Polishing the surface of a metal gives it a highly reflective surface, thus increasing the metallic luster. For e.g., gold, silver and copper. <i>Exception: Lead is a metal with a dull appearance.</i>	Non-metals lack luster. They have a dull appearance. <i>Exception: Crystals of graphite and iodine are lustrous.</i>
<b>4. Density</b>	Metals usually have a high density. For example, the density of gold, mercury and iron is 19.3, 13.6 and 7.6 g/cc., respectively. <i>Exception: Metals such as sodium and potassium are very light as their densities are less than that of water. Lithium is the lightest metal with a density of 0.50g/cc.</i>	Non-metals usually have a low density. <i>Exception: Diamond has a high density, and The density of iodine is higher than that of many metals.</i>
<b>5. Melting and boiling point</b>	Metals usually have high melting and boiling points. <i>Exception: Sodium, potassium, tin and gallium have a low melting point. Mercury has a low boiling and melting point.</i>	Non-metals usually have a low melting and boiling point. <i>Exception: Carbon, silicon and boron possess very high melting and boiling points.</i>
<b>6. Ductility</b>	Metals such as gold, aluminium, copper and silver can be drawn into thin wires very easily without breaking and are therefore said to be ductile. The element tungsten can be drawn into such a fine wire that it is almost invisible. This property of metals is called ductility. <i>Exception : Zinc, mercury and gallium</i>	Non-metals are non-ductile. <i>Exception: Carbon fiber</i>
<b>7. Malleability</b>	Metals are usually malleable because they can be beaten into thin sheets without cracking. Metals such as silver, iron, copper, aluminium and gold expand when stricken with a hammer and are highly malleable. Gold can be beaten into a foil less than 0.0001 mm in thickness. <i>Exception : Zinc and mercury</i>	Non-metals are non-malleable and brittle.



8. <b>Tensile strength</b>	Metals resist breaking when stretched. This is a measure of tensile strength. Metals such as <b>tungsten</b> have a very high tensile strength. <i>Exception: Sodium, calcium, potassium and mercury have low tensile strength.</i>	Non-metals have low tensile strength.
9. <b>Brittleness</b>	Metals are hard but not brittle. <i>Exception: Zinc is brittle at ordinary temperature.</i>	Non-metals are generally brittle.
10. <b>Sonorosity</b>	Most of the metals produce a ringing sound when they are struck.	They are non-sonorous.
11. <b>Conductivity</b>	Metals are good conductors of heat and electricity; <b>silver</b> and <b>copper</b> are very good conductors of heat and electricity. <i>Exception: Bismuth is a poor conductor of heat and electricity.</i>	Non-metals are bad or poor conductors of heat and electricity. <i>Exception: Graphite and gas carbon are good conductors.</i>
12. <b>Solubility</b>	Metals do not generally dissolve in liquid solvents except by chemical reaction. For e.g., magnesium dissolves in hydrochloric acid forming magnesium chloride and hydrogen.	Non-metals dissolve in liquid solvents without any chemical reaction; for e.g., sulphur dissolves in carbon disulphide and iodine in chloroform.
13. <b>Alloy and Amalgam formation</b>	Metals form alloys and amalgams; for e.g., bronze is an alloy of zinc, tin and copper. Similarly, sodium amalgam is an alloy of mercury and sodium.	Non-metals do not form alloys and amalgams. <i>Exception: Carbon, phosphorous, and sulphur can be used in some alloys.</i>

**Table 7.4 Chemical Properties of Metals and Non-Metals**

Chemical properties	Metals	Non-metals
1. <b>Electrochemical nature</b>	Metals are electropositive, i.e. they lose electrons to form positive ions; for e.g., $\text{Na}^+$ , $\text{K}^+$ , $\text{Ca}^{2+}$ , etc. $\text{Na} \rightarrow \text{Na}^+ + \text{e}$ $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}$	Non-metals are electronegative, i.e. they gain electrons to form negative ions; for e.g., $\text{Cl}^-$ , $\text{S}^{2-}$ , $\text{N}^{3-}$ , etc., $\text{Cl} + \text{e} \rightarrow \text{Cl}^-$ $\text{O} + 2\text{e} \rightarrow \text{O}^{2-}$
2. <b>Nature of oxides</b>	Oxides of metals are basic in nature, i.e. the oxides of metals react with water to give bases or alkalis. $\text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{NaOH}$ <i>Exception:</i> Some oxides are amphoteric, such as $\text{Al}_2\text{O}_3$ and $\text{ZnO}$ . Some metal oxides, containing the metal in the higher valence state, are acidic; for e.g., $\text{CrO}_3$ and $\text{Mn}_2\text{O}_7$ .	Oxides of non-metals are acidic in nature, i.e., the oxides of non-metals react with water to give acids. $\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4$
3. <b>Action of acids</b>	Metals dissolve in dilute acids to produce hydrogen gas. $\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2$ <i>Exception :</i> Pb is higher than H in the activity series; it does not displace $\text{H}_2$ from dilute $\text{HCl}$ or $\text{H}_2\text{SO}_4$ .	Non-metals generally do not dissolve in dilute acids.

<b>4. Nature of hydrides</b>	Metals in general do not combine with hydrogen. <i>Exceptions : Some metals (Na, Ca, Li etc.) that combine with hydrogen to form non-volatile unstable hydrides.</i>	Non-metals combine with hydrogen to form stable hydrides; For e.g., NH <sub>3</sub> , H <sub>2</sub> S, PH <sub>3</sub>
<b>5. Nature of chlorides</b>	Metallic chlorides are generally not hydrolysed by water or are only partially hydrolysed. $\text{NaCl} + \text{H}_2\text{O} \rightarrow$ No hydrolysis $\text{AlCl}_3 + 3\text{H}_2\text{O} \rightarrow \text{Al(OH)}_3 + 3\text{HCl}$	Chlorides of non-metals are usually hydrolysed by water. $\text{PCl}_3 + 3\text{H}_2\text{O} \rightarrow 3\text{HCl} + \text{H}_3\text{PO}_3$ $\text{SiCl}_4 + 4\text{H}_2\text{O} \rightarrow \text{Si(OH)}_4 + 4\text{HCl}$
<b>6. Oxidising and reducing action</b>	Metals are reducing agents.	Non-metals are oxidising agents, except carbon, which is a reducing agent.
<b>7. Discharge of ions</b>	Metal ions, being positive, migrate to the negative electrode, i.e., the cathode, and get discharged there by taking up electrons from the cathode.	Non-metal ions, being negative, migrate to the positive electrode i.e., the anode, and get discharged there by losing electrons to the anode.
<b>8. Atomicity</b>	They are monoatomic.	They are diatomic, triatomic and tetraatomic. Exception : Noble gases are monoatomic

The word metal comes from the Greek word 'metallon,' which means quarry or to mine or excavate.

Although new metals are being produced and some metals were difficult to isolate in pure form, there were seven metals known to ancient man. These were gold, copper, silver, mercury, lead, tin, and iron.

### 7.3 CORROSION

Iron pieces exposed to the air and moisture is covered by a reddish-brown scale after some time. This phenomenon is known as rusting and the reddish-brown substance that loosely adheres to the iron is called rust. This is known as corrosion of iron. Similar corrosion is seen in many other metals.

#### Causes for Rusting

Rusting takes place in moist air; not in moisture-free air or in air-free moisture.

Corrosion of metals may be due to the chemical or electrochemical reaction of the metals with the substance present in the surrounding environment; for e.g., rusting of iron.

Except gold, platinum and palladium, almost all the metals undergo corrosion. Thus, rusting of iron, tarnishing of silver, development of a green coating

on copper and bronze are examples of corrosion. Since corrosion causes huge damage to bridges, ships, buildings and other articles made of iron, it costs us crores of rupees every year.

The only metal that is a liquid at ordinary room temperature and pressure is mercury. However, other metals melt close to room temperature. For example, you can melt the metal gallium in the palm of your hand.

#### The Effect of Corrosion Based on The Activity Series

Many metals are easily oxidised. The atmosphere contains a powerful oxidising agent : oxygen. Since metals are always in contact with oxygen, they are vulnerable to corrosion.

Corrosion is defined as the *oxidation of metals exposed to the environment.*

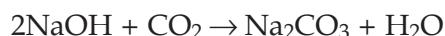
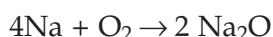
If we go through the activity series of the element, the top-most metals such as sodium and potassium react with air to form oxides.

The tallest free-standing structures in the world are made of metals, primarily the alloy steel. They include the Dubai skyscraper Burj Khalifa, the Tokyo television tower Skytree, and the Shanghai Tower skyscraper.

### Activity Series of Metals:

Activity series
K
Na
Ca
Mg
Al
Zn
Fe
Sn
Pb
H
Cu
Hg
Ag
Au

The elements react with water vapor present in the air to give hydroxide, which absorbs carbon dioxide slowly to form carbonates.



Reactive metals such as sodium and potassium are kept in kerosene oil as they readily react and undergo vigorous reaction with the water vapour present in the atmosphere.

Aluminium and magnesium forms white colour oxides on exposure to the atmosphere as  $\text{Al}_2\text{O}_3$  and  $\text{MgO}$ . Unless a cut or a scratch breaks the oxide layer, the layer prevents further corrosion. Iron forms hydrated ferric oxide, brown in colour, known as rust.

Lead reacts with the moist air to form a thin layer of white coloured oxide, which is a mixture of lead hydroxide and lead carbonate, called basic lead carbonate  $[\text{Pb}(\text{OH})_2]_2\text{PbCO}_3$ .

Copper forms a green deposit on its surface when exposed to moist air. This is usually basic Copper (II) sulphate  $[3\text{Cu}(\text{OH})_2\text{CuSO}_4]$ .

Silver gets tarnished when exposed to the atmosphere, which contains the pollutant  $\text{H}_2\text{S}$  and forms a black coating of  $\text{Ag}_2\text{S}$ .

The noble metals such as gold and platinum do not react with moisture.

### Corrosion of Iron

Corrosion of iron is known as rusting. It is a slow oxidation process of iron with atmospheric oxygen in the presence of water.

Iron (III) hydroxide readily breaks down to form hydrated iron (III) oxide. Rust is represented as  $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$ . The "x" signifies a variable number of water molecules per formula unit.



### Experiment to show that rusting takes place in the following conditions :

Immerse a few clean iron nails in tap water in a test tube and name it as A.

Boil the tap water in a test tube for some time to remove the dissolved air and name it as B. Now, place a few clean iron nails in the test tube, making sure they are completely immersed. Boil the water again and then add a drop or two of oil, which forms a thin layer over the water and blocks the contact of air. Cork the test tube.

Place some anhydrous  $\text{CaCl}_2$  in a test tube and slip a perforated cardboard disk so that it rests a little above the  $\text{CaCl}_2$ . Place some clean iron nails, cork the test tube and name it as C. Anhydrous  $\text{CaCl}_2$ , being a good absorber of moisture, keeps the air inside the test tube moisture-free.

After a few days, we can observe that rusting has taken place only in sample A. This shows that both air and water are necessary for rusting. Though there is a lot of water, the nails in sample B do not rust because they are not in contact with air. The nails in test tube C do not rust because there is no moisture, though there is air.

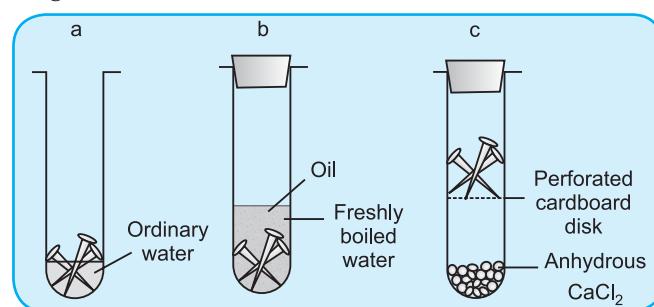


Fig. 7.1 : Corrosion of Iron

### Factors Favourable for Increased Corrosion

Rusting of iron depends upon the following conditions:

1. Impurities in iron, electrolytes, air and moisture, acid-forming gases, such as  $\text{CO}_2$  and  $\text{SO}_2$  lead to rapid rusting during rainy days.
2. Metals placed higher in the activity series corrode more easily.
3. Salt water increases the rate of the corrosion.
4. The presence of pollutants in the atmosphere increases rusting.

### Methods of Preventing Corrosion

Corrosion, especially the corrosion of iron, can be very destructive. Corrosion of metals can be prevented if the metal surface is not allowed to come in contact with moisture, oxygen and carbon dioxide. This can be achieved by the following methods:

#### Barrier Protection

##### Coating with Paints

The simplest method of preventing corrosion is to paint the metal surface. Paint keeps it out of contact with air, moisture, etc., until the paint layer develops cracks.

##### Coating with Oils and Greases

By applying a film of oil and grease on the surface of the iron tools and machinery, the rusting of iron can be prevented since it keeps the metal surface away from moisture, oxygen and carbon dioxide.

##### Galvanizing

Iron is covered with a protective layer of zinc in a process known as galvanizing. Galvanized iron is often used to make metal buckets and chain-link fences. Galvanizing protects iron in two ways. First, the zinc acts as a protective layer. If this layer is broken, the iron is exposed to air and water. When this happens, a protective, invisible, thin layer of basic zinc carbonate  $ZnCO_3 \cdot Zn(OH)_2$  is formed; thus, the iron is still protected without losing its luster.

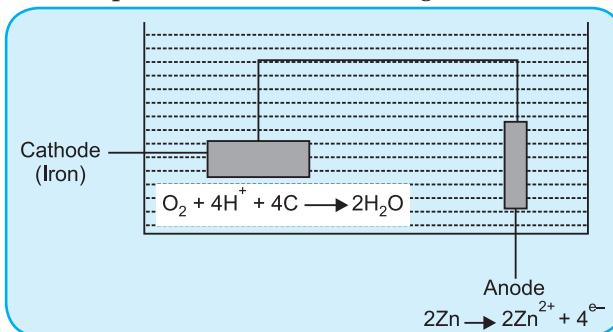


Fig. 7.2 : Galvanization

##### Cathode Protection

It is a process of depositing a stable metal (gold, silver) over a base metal (copper, iron). Iron can be coated with copper by electrodeposition from a solution of copper sulphate. Spoons, handles of bicycles, taps and many other articles are electroplated. Many gift articles, medals, etc., are gold plated or silver plated electrolytically.

#### Advantages of Corrosion

The coated zinc on the iron sheet prevents the rusting of the layer on the surface of iron. When it comes in contact with moisture, oxygen and carbon

dioxide in the air; a protective invisible thin layer of basic zinc carbonate  $ZnCO_3 \cdot Zn(OH)_2$  is formed. Hence, the galvanized iron sheets do not lose their luster and also tends to prevent further corrosion. Galvanized iron is used to make buckets, boxes, utensils etc.

### 7.4 EXTRACTION OF METALS BASED ON THE ACTIVITY SERIES:

**Occurrence of Metals :** Metals occur in nature in various forms. Most of them are found in the *combined state* either in the earth's crust or in oceans as oxide ores, halide ores, sulphide ores, carbonate ores, sulphate ores, etc. However, some of the less reactive metals occur in the *native or free state* because they have little or no affinity for natural substances such as air, water, carbon dioxide etc.

Since the activity series lists the metals in decreasing order of reactivity, it is clear that metals *at the top of the reactivity series* are extremely reactive and exist only in the combined state. For example, sodium and potassium occur mainly as their chlorides ore; calcium and magnesium as carbonates and aluminium as its oxide ore.

Metals in *the middle of the reactivity series* are moderately reactive. For example, zinc, iron, lead and copper usually exist in the earth's crust as oxides, carbonates and sulphides.

Table 7.5 Metals and Their Ores

Metals	Common forms of occurrence	Main ores
K	Chloride Sulphate	Sylvite ( $KCl$ ) Langbeinite ( $K_2SO_4 \cdot 2MgSO_4$ )
Na	Chloride	Sodium chloride ( $NaCl$ )
Ca	Silicate	Calsil ( $2CaO \cdot SiO_2$ )
Mg	Carbonates	Dolomite $CaMg(CO_3)_2$ Magnisite $MgCO_3$
Al	Oxide Fluoride	Bauxite: $Al_2O_3 \cdot 2H_2O$ Cryolite: $AlF_3 \cdot 3NaF$ or $Na_3AlF_6$
Zn	Oxide Sulphide Carbonate	Zincite: $ZnO$ Zinc blende: $ZnS$ Calamine: $ZnCO_3$
Fe	Oxide Sulphide Carbonate	Red Haematite: $Fe_2O_3$ Limonite: $2Fe_2O_3 \cdot 3H_2O$ Magnetite: $Fe_3O_4$  Iron pyrites: $FeS_2$  Siderite or spathic iron: $FeCO_3$

Pb	Sulphide	Galena: PbS
Cu	Sulphide	Copper pyrites: CuFeS <sub>2</sub>
Hg	Sulphide	Cinnabar: HgS
Ag	Chlorides	Horn silver: AgCl
	Sulphide	Argentite: Ag <sub>2</sub> S

Metals of low chemical reactivity (*near the bottom of reactivity series*) exist in nature as either free metals or in a combined state. For example, copper and silver exist as sulphide ores and can also exist in a free state; gold and platinum exist only in the native or free state.

## 7.5 SOME COMMON TERMS USED IN EXTRACTION

### (i) Metallurgy

The process of extracting metals in their pure form from their ores by physical or chemical means is referred to as *metallurgy*.

*Metallurgy is the science and technology of metals. It deals with the production of metals, their purification and alloying.*

### (ii) Minerals

The naturally occurring solid deposits in the earth's crust, which contain metals in the combined state, along with associated earthly impurities such as soil, sand, rocks, etc., are known as minerals.

### (iii) Ores

The minerals from which the metals can be conveniently extracted commercially at low cost and with minimum effort are referred to as ores.

### (iv) Matrix or Gangue

The unwanted earthly impurities such as mud, stones, sand, silica (SiO<sub>2</sub>) etc. associated with the ore are called matrix or gangue.

*For example, calcium silicate is the gangue in the haematite ore and ferrous oxide is the gangue in copper pyrite ore.*

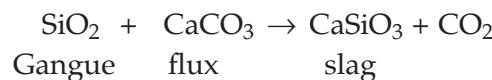
### (v) Flux

A flux is a substance, which is added to molten metals or charged in a furnace to bond with impurities or gangue that can then be readily removed.

*Flux combines with the gangue to form a fusible mass called slag.*

### (vi) Slag

It is the fusible substance obtained by the reaction of flux and impurities (gangue) during the extraction of metals.



*Slag often consists of metallic silicates such as iron silicate and calcium silicate.*

### (vii) Smelting

It is the process of reducing the roasted oxide ore and removing the gangue with the help of an appropriate flux added with the ore. It reduces the oxide ore to metals in the molten condition.

## 7.6 STEPS INVOLVED IN EXTRACTION

Extraction of a metal from its ore involves the following steps:

### Crushing and Grinding

*The ore is crushed into a fine powder in big jaw crushers and ball mills. This process is called pulverization.*

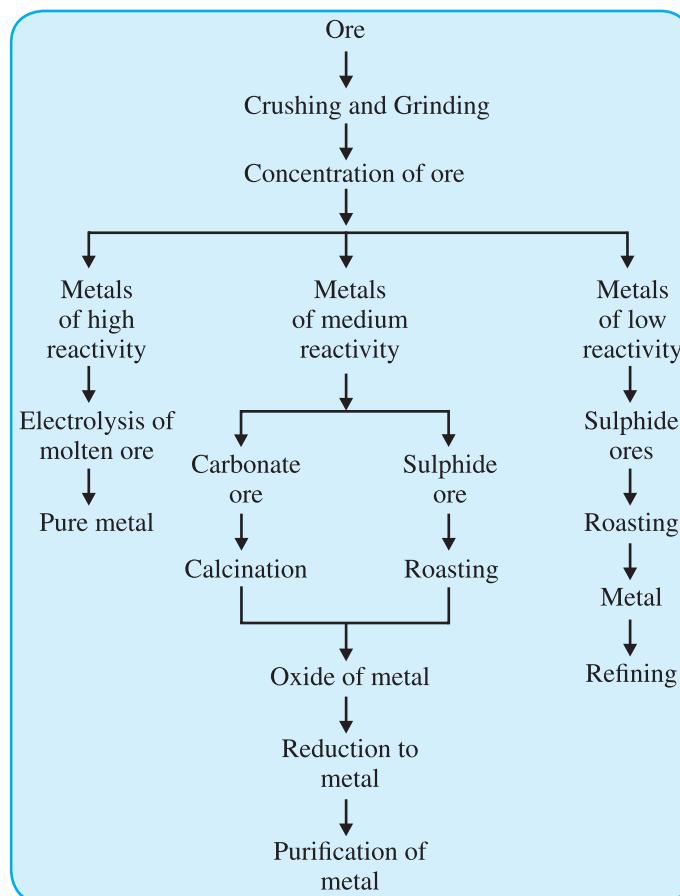
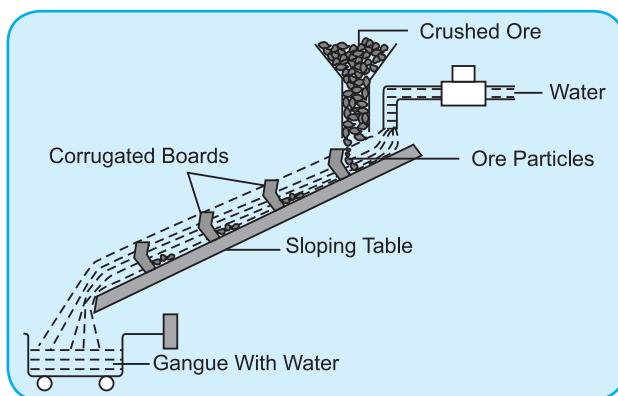


Fig. 7.3 : Steps involved in the Extraction of Metals from their Ores

### Concentration of Ore

*The process of removal of unwanted impurities (or gangue), such as rocks, sand and grit, from the ore is called concentration or ore-dressing and the purified ore thus obtained is called concentrated ore. This process increases the concentration of metal in a mineral ore by the removal of gangue.*

Concentration is a necessary and important step in metallurgical operations. Depending on the type of ore and gangue, various methods are employed for this operation. Some of the methods used for the concentration of ores are discussed below :



**Fig. 7.4 : Gravity Separation :**

Oxide Ores of Iron and Tin (e.g. Haematite and Stannous Oxide) are Generally Concentrated by this Method.

#### (i) Hydraulic Washing (Or) Gravity Separation Method

##### Principle

Differences in the densities of the metallic ore and the gangue particles.

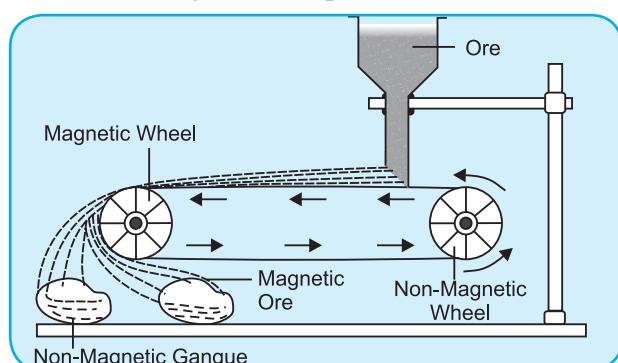
##### Method

This method is used for those ores that are much heavier than the gangue particles present within them. The crushed ore is washed against a stream of water on a vibrating, grooved, sloping table. As a result, the lighter gangue particles are washed away with a stream of running water while the heavier ore particles settle down and remain behind in the grooves.

##### Electromagnetic Separation

##### Principle

A magnet separates the magnetic ore particles from the non-magnetic ore particles.



**Fig. 7.5 : Magnetic Separation : Magnetic Ores of Iron are Concentrated by this Method.**

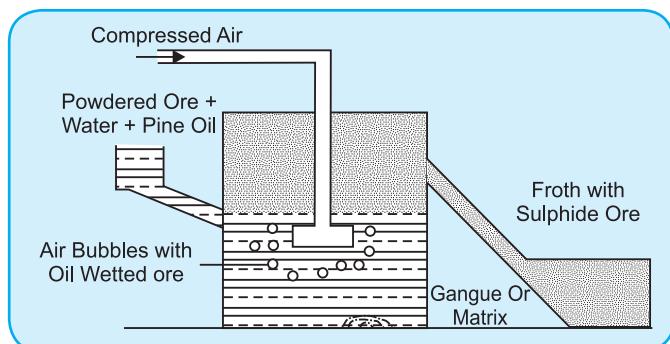
##### Method

This method is useful for those ores that are magnetic in nature. The powdered ore is dropped over the moving conveyor belt, which passes over the two rollers, one of which has magnet in the centre. As the magnetic component of the ore particles come over the roller, they get attracted by the magnet and fall near the roller forming a heap. The gangue particles, being non-magnetic in nature, do not get attracted to the roller and form a separate heap because they fall farther apart due to the high speed of the moving belt.

#### Froth Flotation Process

##### Principle

Differences in the wetting characteristics of the ore and gangue with pine oil and water are used to extract the ore.



**Fig. 7.6 : Froth Flotation Process : Sulphide Ores (For e.g. ZnS, PbS) are Generally Concentrated by this Method.**

##### Method

The crushed ore is taken in a large tank to which water and some quantity of pine or eucalyptus oil is added. The whole mixture is then vigorously agitated with a current of compressed air resulting in the formation of froth. The metallic ore particles stick to the froth and rise to the surface of the tank as scum while gangue particles settle down at the bottom of the tank because they are wetted by water. After some time, the froth containing the metallic ore particles is skimmed off and dried.

#### (iii) Chemical Method or Leaching

##### Principle

Differences in the chemical affinity between the ore and the gangue particles is the main criterion.

The ore is treated with acid, base or some other reagents, so that the ore dissolves in it and the gangue particles remain undissolved. They are then separated by filtration. The solution of the filtrate i.e., the ore, thus obtained, is treated with a suitable

reagent, which precipitates the ore (as in the case of the extraction of aluminium).

### 7.7 ROASTING AND CALCINATION

*Roasting is a process of strongly heating the concentrated ore to a high temperature in excess of air to convert it into a metallic oxide.*

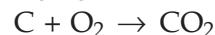
*Calcination is the process of the conversion of concentrated ore into metal oxide (oxidation) by strong heating in the absence of air at a temperature insufficient to melt the ore.*

**Table 7.6 Roasting and Calcination**

Roasting	Calcination
(i) Ore is heated in an excess of air.	(i) Ore is heated in absence of air.
(ii) Sulphide ores are generally roasted. Hence, sulphur dioxide gas is emitted. $2\text{ZnS} + 3\text{O}_2 \rightarrow 2\text{ZnO} + 2\text{SO}_2 \uparrow$ $4\text{FeS}_2 + 11\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3 + 8\text{SO}_2 \uparrow$	(ii) Carbonates and hydrated ores are generally calcinated. Hence, carbon dioxide gas and water are given off. $\text{ZnCO}_3 \rightarrow \text{ZnO} + \text{CO}_2$ $\text{CuCO}_3 \cdot \text{Cu(OH)}_2 \rightarrow 2\text{CuO} + \text{H}_2\text{O} + \text{CO}_2 \uparrow$
(iii) Volatile impurities are removed as oxides ( $\text{SO}_2$ , $\text{P}_2\text{O}_5$ , $\text{As}_2\text{O}_3$ ) and the ore becomes porous and more reactive.	(iii) Moisture and organic impurities are removed.

#### Reduction of metal oxide to free metal

We have learnt that the oxides of metals below Al in the activity series can be reduced to the respective metals on being heated with C. Apart from C, there is a possibility of CO acting as the reducing agent – CO is a stronger reducing agent than C :

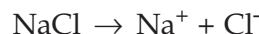


Many metal oxides, such as  $\text{FeO}$ ,  $\text{Fe}_2\text{O}_3$ ,  $\text{PbO}$  and  $\text{CuO}$  are easily reduced by C, CO and  $\text{H}_2$  at higher temperatures to the corresponding metals. However, many other metal oxides, such as  $\text{Na}_2\text{O}$ ,  $\text{CaO}$ ,  $\text{MgO}$  and  $\text{Al}_2\text{O}_3$  are not reduced easily by these reducing agents.

If the metal is highly active, it's quite difficult to reduce its oxide from the activity series. We can understand that moving down the activity series; the reduction of the oxide increases.

The oxides of highly active metals such as potassium, sodium, calcium, magnesium and

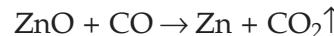
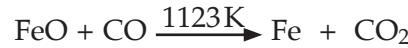
aluminium have a greater affinity towards oxygen and so cannot be reduced by carbon, carbon monoxide or hydrogen. These metals are obtained by electrolytic reduction of fused metallic salts (halides or oxides).



The metal oxides are then reduced to the corresponding metals by using suitable reducing agents.

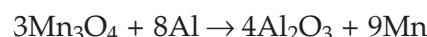
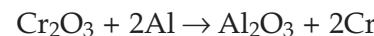
#### Carbon as Reducing Agent

The oxides of moderately reactive metals, such as iron, zinc, etc., are reduced to the metal by heating with carbon in the form of coal, coke, charcoal and carbon monoxide.



#### Aluminium as Reducing Agent

Here, aluminium is used as the reducing agent because it is a highly reactive metal, which can displace metals of lower reactivity from their compounds. Hence, the oxides of chromium and manganese are reduced by an alumino thermic process or GOLD SCHMIDT THERMAL process.



#### Refining of Metals

The metals obtained by various reduction processes contain certain impurities such as :

(i) presence of other metals, non-metals such as silicon or phosphorus

(ii) unreduced oxides and sulphides of the metal

The impure metal is again subjected to certain purifying processes, which depend upon the metal. The process of purification of impure metals by removing metallic and non-metallic impurities is known as refining of metals.

The refining process depends on the following factors :

(i) nature of the metal

(ii) nature of impurities

(iii) purpose for which the metal is to be used

#### 1. Distillation

This process is employed for purification of volatile metals such as mercury, zinc, and cadmium, which contain non-volatile impurities. The impure metal

is heated in a retort, and its vapors are separately condensed in a receiver. While the pure metal distills over, the non-volatile impurities are left behind in the retort.

### 2. Liquation

This method is used for refining metals having low melting point such as tin, lead etc. The impure metal is gently heated on the sloping hearth of a furnace. The metal melts and drains away leaving behind the infusible materials on the hearth.

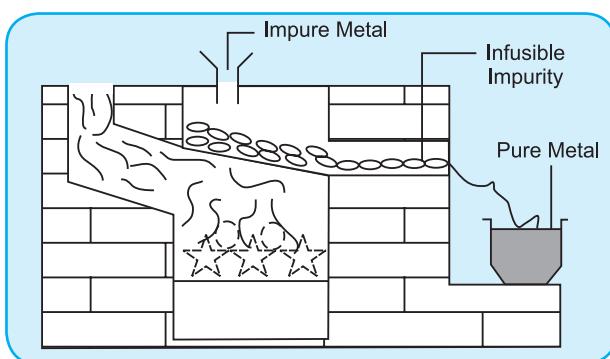


Fig. 7.7 : Liquation

### 3. Electro-refining

This is the most economical and widely used method for refining impure metals, such as copper, zinc, tin, nickel etc.

The impure metal is made the **anode**, while the thin plate of the pure metal is the **cathode**. The electrolyte used is a salt solution containing the metallic radical common to the metal that has to be refined.

On passing the electric current, the pure metal deposits at the cathode and impurities from the anode settle down at the bottom of tank as *anode mud*.

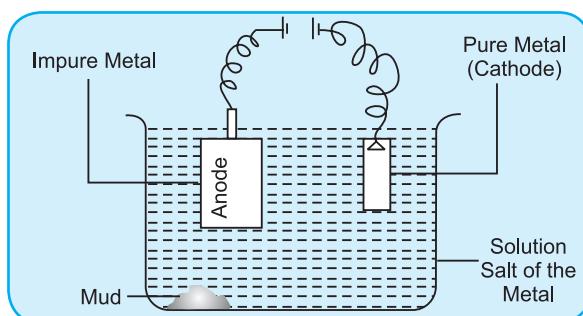


Fig. 7.8 : Electro-refining

### 4. Oxidation

This method is used to refine metals containing volatile impurities, which are more readily oxidised than the metal itself. The impure metal is melted and exposed to air in a suitable furnace. The oxides of impurities rise to the surface and are removed while the pure, molten metal remains behind.

The volatile impurities such as phosphorous, carbon, sulphur and arsenic are removed by this method.

### Common ores of Iron and Zinc

Ore	Chemical name and formula
Haematite	$\text{Fe}_2\text{O}_3$
Limonite	$\text{Fe}_2\text{O}_3 \cdot \text{H}_2\text{O}$
Magnetite	$\text{Fe}_3\text{O}_4$

Ore	Chemical name and formula
Zincite	$\text{ZnO}$
Zinc blende	$\text{ZnS}$
Calamine	$\text{ZnCO}_3$

## 7.8 ALUMINIUM

### Common Ores of Aluminium

Ore	Chemical name and formula
Bauxite	Hydrated aluminium oxide, $\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$
Cryolite	Sodium aluminium fluoride, $\text{Na}_3\text{AlF}_6$
Corundum	Anhydrous aluminium oxide, $\text{Al}_2\text{O}_3$

### Aluminium

Symbol	Al
Colour	Silvery white
Atomic number	13
Atomic mass	27
Specific gravity	2.2
Electronic configuration	2,8,3
Valency	3
Position in periodic table	Period 3, Group 13 (III A)

### Discovery and Occurrence

Aluminium was discovered by *Oersted* in 1825. Later in 1886, Charles Martin Hall discovered the process of extracting aluminium from bauxite by electrolysis.

It is the most abundant metal constituting nearly 8% of the earth's crust. It is always chemically combined with other elements and therefore, never occurs freely in nature.

### Extraction of Aluminium from Its Ore

The extraction of aluminium requires the following stages :

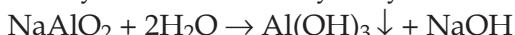
1. Purification of bauxite
2. Electrolytic reduction

### Purification of bauxite by Baeyer's process

The bauxite ore is powdered and roasted to convert any ferrous oxide impurity into ferric oxide. Then, it is mixed with a concentrated solution of caustic soda ( $\text{NaOH}$ ) solution for 2 to 8 hours at  $140^\circ\text{C}$  to  $150^\circ\text{C}$  in an autoclave under pressure. Aluminium oxide present in the ore dissolves to give sodium meta aluminate, while the impurities (known as red mud) are left behind.



Ferric oxide and silica impurities are removed by filtration. The sodium meta aluminate solution is diluted with water and allowed to attain  $50^\circ\text{C}$ ; it is then agitated for a long time. Freshly prepared aluminium hydroxide is added to act as a seeding agent. The whole of sodium meta aluminate precipitates as aluminium hydroxide due to hydrolysis.

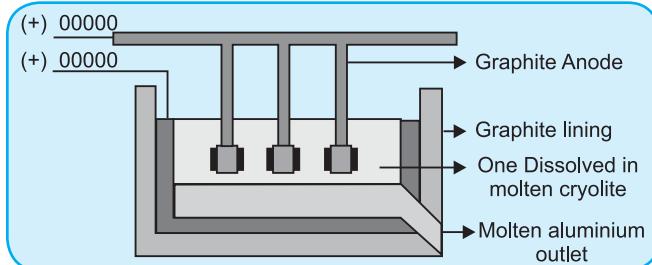


The precipitate of aluminium hydroxide is filtered, washed and ignited at  $1000^\circ\text{C}$  (i.e., strongly heated) to form pure alumina.



### Electrolytic reduction of alumina by Hall's electrolytic process

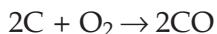
- Pure alumina is dissolved in molten cryolite and mixed with a little amount of **fluorspar**, which lowers the melting point of electrolyte.
- The molten electrolyte is taken in a rectangular cast iron tank lined with a layer of graphite and at the bottom of the tank, there is an outlet where the molten aluminium can be driven out.
- The iron tank lined inside with graphite acts as the cathode.
- A bunch of graphite rods is suspended in the molten electrolyte to act as the anode.
- The electrolysis is carried out at  $900\text{-}950^\circ\text{C}$ . During the electrolysis, molten aluminium is produced at the cathode and oxygen gas is evolved at the anode.



**Fig. 7.9 : Aluminium Extraction**



The oxygen gas evolved at the anode reacts with graphite anodes and forms carbon monoxide which further forms carbon dioxide gas.



The anodes, thus burn away. Therefore, they must be replaced from time to time.

When the supply of alumina decreases, the voltage suddenly jumps from 40 to 60 volts. If a bulb is attached to the apparatus, it will glow brightly; then, the fresh alumina is added. Aluminium obtained is 99.8% pure. Further purification is done by electrolysis.

#### NOTE

**Cryolite :** Lowers the fusion temperature from  $2050^\circ\text{C}$  to  $950^\circ\text{C}$  and enhances conductivity.

**Fluorspar And Cryolite :** They act as a solvent for the electrolytic mixture. They increase its conductivity since pure alumina is almost a non-conductor of electricity.

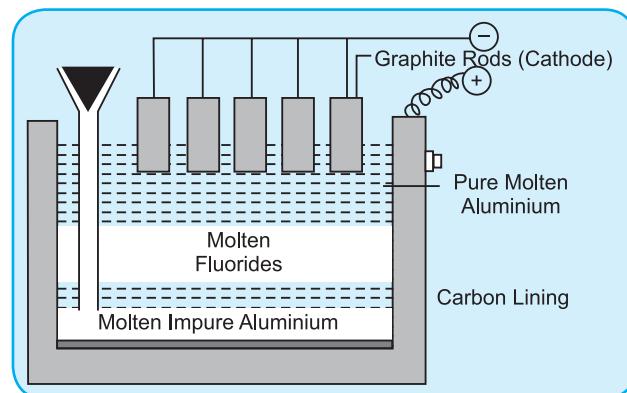
**Powdered Coke :** It is sprinkled over the surface of the electrolytic mixture because it reduces heat loss by radiation. It prevents the burning of the anode.

### Refining of aluminium (Hooper's electrolytic process)

The process is done based on the different specific gravities of the molten substances. The molten impure aluminium is formed at the bottom layer. The bottom layer has a carbon lining and serves as anode. Pure molten aluminium with carbon electrodes serves as the cathode in the top layer.

A mixture of fluorides of sodium, barium and aluminium are placed in the middle layer.

When current is passed, aluminium from the middle layer passes to the top layer and an equivalent amount of it shifts from the lower layer to the middle one. Pure aluminium is tapped from the top.



**Fig. 7.10 : Aluminium Refining**

**Anode:** Al (impure) – 3e → Al<sup>3+</sup>

**Cathode:** Al<sup>3+</sup> + 3e<sup>-</sup> → Al (pure)

## 7.9 PROPERTIES OF ALUMINIUM

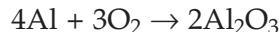
### Physical Properties

- It is a silvery white metal with high tensile strength.
- It is extremely light with a specific density of 2.7g/cm<sup>3</sup> at 20°C.
- It is malleable and ductile.
- It melts at 660°C and boils at 2050°C.
- It is a very good conductor of heat and electricity.

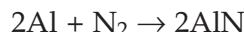
### Chemical properties

#### 1. Reaction with Air

Aluminium gets oxidized to oxide at room temperature and forms a thin, tough protective coating of aluminium oxide on its surface that protects it from further attack of air. *This is the reason that aluminium undergoes minimum corrosion on exposure to air.*

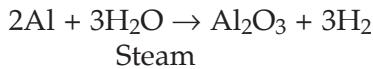


Powdered aluminium burns in air at about 800°C with a bright light forming the metal oxide and nitride.



#### 2. Reaction with water

Hot and cold water has no action on aluminium. However, when steam is passed over heated aluminium, aluminium oxide is formed with the liberation of hydrogen gas.

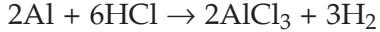


#### 3. Amphoteric Nature

It reacts with strong bases as well as acids; therefore, it is amphoteric in nature.

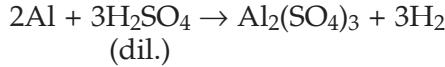
#### Reaction with Acids

Aluminium reacts with dil. and conc. HCl to form aluminium chloride and hydrogen gas.

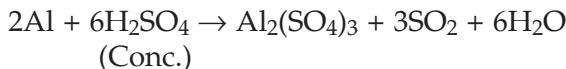


Initially, the reaction is slow due to the protective layer of aluminium oxide Al<sub>2</sub>O<sub>3</sub> present in the surface of the aluminium which resists the attack of the acid.

Dil. sulphuric acid reacts with aluminium to liberate hydrogen.



Aluminium reacts with hot concentrated sulphuric acid to form aluminium sulphate and sulphur dioxide.



### NOTE

Dil. or conc. nitric acid does not attack the aluminium metal because it renders aluminium passage due to the formation of an oxide film on its surface.

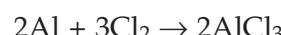
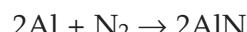
#### Reaction with Alkalies

Aluminium forms aluminates and liberates hydrogen on reaction with boiling aqueous NaOH and KOH.



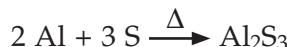
#### 4. Action on Non-Metals

Heated aluminium combines directly with non-metals.



#### 5. Action of aluminium on sulphur

Aluminium sulfide is readily prepared by ignition of the elements



The product will be created in a fused form; it reaches a temperature greater than 1100°C and may melt its way through steel. The cooled product is very hard.

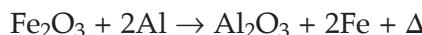
#### 6. Action of aluminium on carbon

When aluminium reacts with carbon in an electric furnace, aluminium carbide is formed.



#### 7. Reducing Action

Aluminium is a powerful reducing agent; therefore, when a mixture of aluminium powder and iron oxide is ignited, the latter is reduced to metal. This process is known as aluminothermy.



## 7.10 USES OF ALUMINIUM

- Aluminium is used in making electric cables and transmission wires because of its high electrical conductivity.
- It is used in utensils because of its high thermal conductivity.
- Aluminium alloys are used in aeroplane parts, surgical instruments etc.; the alloys of aluminium



are known for their lightness and lustrous properties.

- It is used as a reducing agent in the extraction of chromium and manganese by Goldschmidt alumino-thermic process.
- Aluminium powder mixed with a suitable linseed oil is used to prevent corrosion of iron poles.
- Since it is not attacked by concentrated nitric acid, it is used in chemical plants and for transporting nitric acid.
- Aluminium powder is used in making silver paints, fireworks, flash light powders and thermic welding.
- A mixture of aluminium powder and ammonium nitrate called ammonal is used in explosives.
- It is used in building and construction works.
- It is used in the form of thin foils for wrapping chocolates, medicines, etc., and in capping milk bottles, in making tubes for tooth-paste, etc.
- It is used for making flash bulbs in photography.
- Being an excellent reflector of light, aluminium is used for making mirrors for reflecting telescopes.
- Aluminium, being a good conductor of electricity, is used for the manufacture of cables for power transmission.
- As aluminium is unaffected by seawater, ships are made of alloys of aluminium.

## 7.11 FORMS AND USES OF IRON

Depending upon the carbon content, various forms of iron are known such as :

### (a) Cast iron or pig iron

It is the most impure form of iron containing about 2.5 – 5 % of carbon, along with impurities such as sulphur, silicon, phosphorus, manganese, etc. It is very hard, brittle and corrosion resistant.

**Use :** It is used in sewage pipes, gutter covers, weights and railings.

### (b) Wrought iron

It contains carbon and other impurities less than 0.1 - 0.25 % of content. It is malleable and can be easily welded.

#### Use

It is used in making wires, chains, horse shoes and electromagnets.

### (c) Steel

Steel is an alloy of iron and carbon with 0.25% - 2% of carbon content. It is highly malleable and ductile.

#### Use

It is used in the construction of machines, automobiles, buildings, instruments, etc. and also in various alloys.

The properties of steel can be changed by heat treatment called *tempering*.

When the red-hot steel (800°C) is immersed immediately into cold water or oil, it becomes harder and more brittle, like glass. This process is called *quenching of steel*; specific machine parts are made by quenching.

When the quenched steel is reheated and regulated to a temperature of 220°C - 330°C and then allowed to cool slowly, it retains its hardness but its brittleness vanishes. This is because the very slow cooling of the metal relieves the strain set up during heat treatment.

This process is known as *annealing of steel*. Razor blades, springs, etc. are made by annealing.

### Different Types of Steel

Based on carbon content, there are three types of steel.

#### (i) Mild Steel

It has the least carbon content (0.1-0.15%). It is used for making wires, pipes, sheets, etc.

#### (ii) Medium Steel

It contains 0.2 to 0.5% carbon. It is harder than mild steel and is used for constructing rails, wheels, boilers etc.

#### (iii) Hard Steel

It contains a higher percentage of carbon (0.5 - 1.5%) when compared to the other two types of steel. It is very hard and is used for making machine parts.

## 7.12 USES OF ZINC

Zinc has a strong electropositive character because of its tendency to lose electrons. It is, therefore, used for the following purposes :

1. About 23% of all zinc produced is used for coating iron to prevent rusting. This process is called *galvanization*. This coating is done mechanically either by dipping the hot iron sheets into molten zinc at 300°C or by spraying molten zinc on the

surface of iron articles or by electrolysis. Sheets of galvanized iron are used for making pipes, buckets, dustbins, bath tubs, water tanks, wire and roofs.

2. Zinc readily forms  $Zn^{2+}$  ions because of its strong electropositive nature. It is, therefore, used for making *dry cell containers*, which serve as the cathode or negative electrode of the cell.
3. Zinc granules (zinc + copper) are used in laboratories.
4. It is used for preparing lithographic plates for printing.
5. Zinc, in the form of zinc dust, is used as a reducing agent in various industries for manufacturing drugs, dyes, perfumes, etc.
6. Various compounds of zinc are used in paints, electroplast (antiseptic), preservatives for leather and mordant for textiles. Zinc carbonate is used especially for making cosmetics, such as face creams, lotions, etc.
7. Alloys of zinc such as *brass*, *bronze* and *german silver* are used extensively.

Alloys of Zinc	Uses
Brass	Musical instruments, electrical fittings, etc.
Bronze	Statues
German silver	Jewellery and coins

### 7.13 ALLOYS AND AMALGAM

Transition metals form alloys with each other. This is because they have almost a similar size and the atoms of one metal can easily take up positions in the crystal lattice of the other.

A *homogeneous mixture*, i.e., a solid solution of a metal with other metals or non-metals is an alloy.

We cannot make gold ornaments without mixing copper. An alloy melting in the range of about  $51^{\circ}\text{C}$  to  $260^{\circ}\text{C}$  usually contains bismuth, lead, tin, etc. These alloys are called fusible alloys.

#### Example :

Brass is an alloy of zinc and copper.

Steel containing up to 10% of elements, such as chromium, molybdenum, nickel, etc., usually with a low percentage of carbon is known as an alloy steel. These added elements improve hardness, wear resistance, toughness and other properties.

An alloy of mercury with another metal that is solid or liquid at room temperature according to the proportion of mercury present is an amalgam.

All metals form alloys or amalgams with mercury, except iron. Amalgams are used as dental materials, in the concentration of gold and silver from their ores, and as electrodes in various industrial and laboratory electrolytic processes.

#### For Example :

- An alloy of mercury and silver is used in dental fillings.
- An alloy of mercury and tin is used in silvering mirrors.

Amalgam used in dental work requires the following composition: silver 65%, copper 6%, zinc 2% and tin 25%. They are generally more useful than the single metals for the following reasons:

- An alloy is usually harder and tougher than the metals it is made of. For example, steel and duralumin are stronger than iron and aluminium, respectively.
- An alloy resists corrosion.

#### Purpose of Making Alloys

Alloys are more useful than pure metals because a pure metal has a few important properties such as melting point, boiling point, density, specific gravity, high malleability, ductility etc. These properties can be changed and modified and enhanced by alloying them with some other metal or non-metal according to our need.

#### Reasons for Alloying:

The reasons for making alloys is to lower the melting point, to enhance tensile strength, corrosion resistance, hardness of a metal, to modify the colour and to provide better castability.

The majority of alloys are prepared by mixing metals in the molten state; then, the mixture is poured into metal or sand moulds and allowed to solidify. Generally, the major ingredient is melted first, and then the others are added to it and should completely dissolve.

For example, while soldering, an electrician must melt lead and tin and mix and cast the alloy into a thin wire.

## 7.14 A LIST OF SOME COMMON ALLOYS ALONG WITH THEIR COMPOSITION AND USES.

**Table 7.7 Uses of Alloy and its Composition**

Main Metal	Alloy	Composition	Properties	Uses
Aluminium	Duralumin	Al = 95%, Cu = 4%, Mg + Mn = 1%	Light-weight, strong, resistance to corrosion	Aircraft bodies
	Magnalium	Al = 95%, Mg = 5%	Highly ductile, light, strong and resistant to corrosion	Scientific instruments, utensils and powerful magnets
	Alnico	Fe = 50%, Ni = 20%, Ca = 10%, Al = 10%	Light, shiny and resists corrosion	
Zinc and Copper	Brass	Cu = 60 – 80%, Zn = 20 – 40%	Lustrous, easily cast, malleable, ductile	Utensils, ornaments, door handles, electrical fittings
	Bronze	Cu = 75 – 90%, Sn = 10 – 25% and some Zn	Hard, brittle	Statues, coins and utensils
	German silver	Cu = 50%, Zn = 30%, Ni = 20%	Hard, lustrous	Coins, decorative articles
	Bell metal	Cu = 80%, Sn = 20%	Hard, brittle, sonorous	Bells and gongs
	Gun metal	Cu = 88%, Zn = 1%, Sn = 10%, Pb = 1%	Hard, brittle, easy cast	Rifle, Cannons
Iron	Stainless steel	Cr = 17 – 19%, Ni = 7 – 9%	High tensile strength, lustrous, does not rust	Cutlery, utensils and many general uses
	Manganese steel	Mn = 10 – 15%	High tensile strength	Rails, rock drills
	Tungsten steel	Fe = 79%, Sn = 20%, C = 1%	Heat resistant	More heat generating machine
	Nickel steel	Fe = 97%, Ni = 2%, C = 1%	High toughness	Cables, aircrafts parts
	Invar	Ni = 36%, Fe = 64%	Co-efficient of expansion is very small	Measuring tapes, pendulums
Tin and lead	Solder	Sn = 67%, Pb = 33%	Moisture resistant, low melting point	For soldering wires and electronic components
	Fuse metal	Bi = 50%, Sn = 28%, Pb = 22%	Very low melting point	Electric fuses

### INTEXT QUESTIONS

- What are metals? Give few examples.
- Name the metal and non-metal present in abundance on the earth's crust.
- Define metals and non-metals based on electron loss or gain.
- Name the non-metal that is liquid at ordinary conditions.
- Name the non-metal that is a lustrous solid.
- State the position of the following in the periodic table : (a) alkali metals (b) alkaline earth metals (c) iron and zinc (d) aluminium
- What metallic property is shown by the non-metal graphite?
- How many valence electrons are present in the alkali metals and alkaline earth metals ?
- Give the general characteristics of (a) alkali metals (b) alkaline earth metals with reference to

- (i) bonding (ii) action of air (iii) action of water  
(iv) action of acid.
10. Explain malleability of metals.
  11. Write the atomicity of metals and non-metals.
  12. Name two metals that are liquid at room temperature, two metals that are soft, two metals that react with cold water and a metal that lacks ductility.
  13. What is the chemical composition of rust ?
  14. Name two important methods for preventing the corrosion of iron.
  15. Name a non-metal which is lustrous and conducts electricity.
  16. Explain the activity series accounting for each of the following phenomena :  
a. Occurrence of metals b. Tendency to corrosion  
c. Reaction with water d. Reaction with acid
  17. Give the balanced reactions for the following :  
a. sodium is dropped in water b. magnesium reacts with boiling water c. red hot iron reacts with steam d. iron reacts with dilute HCl.
  18. Give a short account of heating effect on metal carbonates based on the activity series.
  19. Why are alkali metals kept in kerosene oil?
  20. What is the chemical substance which deposits on iron?
  21. What is corrosion? What are necessary conditions for corrosion?
  22. State under what conditions corrosion is faster.
  23. What is rust? Give the equation for the formation of rust.
  24. Which metals do not corrode easily?
  25. Define the term 'metallurgy'.
  26. What are the various steps involved in metallurgy?
  27. Distinguish between a mineral and an ore; an ore and a metallic compound.
  28. Which metals can be extracted from each one of the following ores :  
(a) bauxite (b) calamine (c) Haematite.
  29. State three objectives achieved during the roasting of ores.
  30. Give the principle of :  
(a) hydrolytic method (b) froth floatation  
(c) electromagnetic separation
  31. Explain the terms :  
(a) flux (b) gangue  
(c) slag (d) smelting
  32. Why does iron or zinc not occur freely in nature?
  33. Compare roasting and calcination.
  34. Explain briefly how the activity series helps to study the reduction of metallic oxides.
  35. (a) On which factors does the purification of metals depend?  
(b) Name the methods used for purification.  
(c) How is electro-refining done?
  36. Where is aluminium present in the periodic table ?
  37. Name the principle ore from which aluminium is extracted.
  38. Write the chemical names of aluminium ores with their formulae.
  39. Write three balanced equations for the purification of bauxite by Hall's process.
  40. Name a chemical used for dissolving aluminium oxide. In which state of subdivision is the chemical used ?
  41. Write an equation for the reaction that takes place at the anode during the extraction of aluminium by the electrolytic process ?
  42. Mention one reason for the use of aluminium in thermite welding.
  43. What are the impurities present in bauxite ?
  44. How is pure aluminium oxide converted into a conducting solution?
  45. What is red mud and how is it removed ?
  46. Why is electrolytic reduction done to obtain aluminium ?
  47. How does aluminium react with  
(a) air (b) water (c) acid (d) base?
  48. What is the role of cryolite in the electrolytic reduction of alumina in Hall's process ?
  49. Why does aluminium suffer less corrosion even after being a more active metal than iron ?
  50. Explain and give reasons why aluminium vessels should not be cleaned with powders containing alkalis.
  51. Give the name and formulae of the main ores of iron and zinc.
  52. How is the main ore of aluminium concentrated ?
  53. Why is it that 'the food containing iron salts' should not be cooked in aluminium utensils ?
  54. What are the electrodes used in the extraction of aluminium ?

55. Explain the reasons :

- In the electrolytic reduction of alumina, the graphite anode is gradually consumed.
- Roasting is carried out on sulphide ores and not on carbonate ores.
- Carbon can reduce lead oxide but not aluminium oxide.

56. (a) Why is flux used in the blast furnace ?

(b) What does it form with silica present in the ore ?

(c) How is it removed ?

57. Name an ore, which is concentrated by :

- Froth floatation process
- Magnetic separation

58. Distinguish between electrolytic methods of reduction and refining.



### POINTS TO REMEMBER

- ◆ Metals are lustrous, crystalline solids of high strength, high melting and boiling points.
- ◆ Metals are good conductors of heat and electricity.
- ◆ However, a liquid under ordinary conditions, Hg is a metal.
- ◆ Though electropositive, H is a non-metal. Though lustrous solids, graphite and iodine are non-metals.
- ◆ Non-metals are elements that, in general, are
  - Non-lustrous
  - Bad conductors of heat and electricity
  - Brittle, if solid.
  - Electronegative, i.e., they form negative ions by gaining electrons.
- ◆ All elements on the left-hand side of staircase demarcation in the periodic table are metals. Non-metals are much fewer in number. Including the semi-metals, there are only 24 non-metals – 12 solids, 1 liquid and 11 gases. The oxides of metals down to Al in the activity series are not easily reduced by C, CO and H<sub>2</sub>.
- ◆ Active metals such as K, Ca, Na, Mg, Al, Zn, Fe, and Pb are found in the combined state in nature.
- ◆ Noble metals such as Au and Pt are found in the combined as well as the native state in nature.
- ◆ Minerals are naturally occurring chemical substances, along with earthy materials, found in the earth's crust.
- ◆ An ore is a mineral from which a required metal can be extracted profitably.
- ◆ The impurities present in the ore are known as gangue.
- ◆ A substance added in the furnace during smelting to remove the gangue is known as flux.
- ◆ A substance formed when flux reacts with gangue is called slag.

- ◆ The extraction of a metal consists of three steps:
  - Concentration of the ore
  - Production of the crude metal
  - Purification of the crude metal
- ◆ The concentration of the ore is affected by the following methods :
  - Gravity separation
  - Froth floatation
  - Magnetic separation
  - Leaching
- ◆ The production of a metal involves the following steps.
  - Calcination** : Heating an ore below its fusion point to remove any volatile material is known as calcinations.
  - Roasting** : Heating an ore below its fusion point, in a regular supply of air, to convert the metal into its oxide is known as roasting.
  - Flux** : It is a substance that chemically reacts with the gangue to form a fusible mass known as slag.
- ◆ Aluminium is obtained by the electrolytic reduction of bauxite dissolved in cryolite.
- ◆ The Goldschmidt-Thermite process is used to weld broken rails, iron parts in ships, etc.
- ◆ Corrosion of iron takes place in moist air but not in moisture free air or in air-free moisture.
- ◆ Rusting can be prevented by painting the metal surface, galvanizing or electroplating the metal with Cr or Ni.
- ◆ In galvanization, the iron article is dipped in molten Zn and taken out, thus the surface of iron is coated by a thin, protective layer of Zn.
- ◆ An alloy is a homogeneous mixture of two or more metals or of one or more metals with certain non-metallic elements.

## EXERCISES

## A. FILL IN THE BLANKS

1. Metals are ..... while non-metals are poor conductors of electricity.
2. Metals are malleable while non-metals are .....
3. Metals form positive ions while non-metals form .....
4. Non-metals form acidic oxides while metals form .....
5. The ore from which aluminium is extracted must first be treated with ..... so that pure aluminium oxide can be obtained.
6. Pure aluminium oxide is dissolved in ..... to make a conducting solution.
7. ..... has the maximum metallic character.
8. Aircrafts bodies are made up of .....
9. Rust is chemically called .....
10. Heating an ore below its fusion point in order to remove any volatile matter is known as .....
11. Non-metals are ..... in nature.
12. The valency of an alkali metal is .....
13. The ..... alloy is used in the electric fuse while soldering.
14. ..... is the dental amalgam.
15. Metals are good ..... because they are electron ..... [2016]

## B. MULTIPLE CHOICE TYPE QUESTIONS

Choose the correct answer from the options given below :

1. Transition metals belongs to group  
 (A) 3-12                          (B) II A  
 (C) I A                            (D) 13-17
2. Metal liquid at room temperature :  
 (A) Mercury                      (B) Potassium  
 (C) Sodium                        (D) Gallium
3. Chemical formula of Rust  
 (A)  $\text{Fe}_2\text{O}_3 \cdot \text{xH}_2\text{O}$             (B)  $\text{ZnCO}_3$   
 (C)  $\text{CuSO}_4 \cdot \text{xH}_2\text{O}$                     (D)  $\text{Na}_2\text{O}_3$
4. Formula of Haematite  
 (A)  $\text{Fe}_2\text{O}_3$                         (B)  $\text{Al}_2\text{O}_3$   
 (C)  $\text{ZnCO}_3$                         (D)  $\text{Na}_3\text{AlF}_6$
5. Atomicity of metals is :  
 (A) Mono atomic                    (B) Triatomic  
 (C) Diatomic                        (D) Tetra-atomic

## C. CHOOSE THE MOST APPROPRIATE ANSWER FOR EACH OF THE FOLLOWING

- (i) Among the elements given below, the element with the least electronegativity is :  
 (A) Lithium                        (B) Carbon  
 (C) Boron                           (D) Fluorine [2015]
- (ii) This is not an alloy of copper :  
 (A) Brass                           (B) Bronze  
 (C) Solder                           (D) Duralumin [2015]

## D. MATCH THE COLUMNS

Column I	Column II
1. Alkali metals	Brittle
2. Galena	Group 13-17
3. Iron	$\text{PbS}$
4. Metals	Rusting
5. Non-metals	Sonorous
6. Lanthanides, Actinides	Group IA

## E. STATE ONE RELEVANT OBSERVATION FOR THE FOLLOWING

- (a) At the Anode when aqueous copper sulphate solution is electrolysed during copper electrodes. [2015]
- (i) Differentiate between the terms **strong electrolyte** and **weak electrolyte**. (stating any two differences) with examples. [2015]
- (b) Element X is a metal with a valency 2, Y is a non-metal with a valency 3.  
 (i) Write an equation to show how Y forms an ion.  
 (ii) If Y is a diatomic gas, write an equation for the direct combination of X and Y to form a compound. [2015]

## F. VERY SHORT ANSWERS

1. Give answer for the following :
  - (i) Metal liquid at room temperature.
  - (ii) Non-metal that is a conductor of electricity.
  - (iii) Neutral oxide.
  - (iv) Metallic oxide that cannot be reduced by hydrogen.
  - (v) Non-metal that has luster.
2. What is the name of the most common ore of iron and what is its chemical formula?
3. How is zinc vapour condensed?

4. Arrange the metals in the decreasing order of reactivity.
5. What is the composition of stainless steel?
6. Name a non-metal, which is :
- A liquid at ordinary conditions
  - A good conductor of electricity [ICSE 2009]
  - A lustrous solid
  - Electropositive
7. What metallic property does the non-metal graphite show?
8. Among K, Cu and S, choose the element:
- Which forms positive ions most easily
  - Which reacts most vigorously with water to liberate H<sub>2</sub>
  - The oxide of which is easily reduced by H<sub>2</sub>
  - Which forms an acidic oxide
9. How many valence electrons are present in
- alkali metals
  - alkaline earth metals?
10. Among Cu, Fe, Mg, Na and Zn, select a different metal in each case, which
- Does not react with dilute HCl
  - Forms 2<sup>+</sup> and 3<sup>+</sup> ions
  - Has a hydroxide that reacts with both acids and alkalis
  - Does not react with cold water but reacts with steam when heated
11. What is the common feature of the electronic configuration of the alkali metals?
12. What do you mean by the term ore?
13. What is the term used for the impurities associated with an ore?
14. Name the main constituent metal in the following alloys: [ICSE 2010]
- Duralumin
  - Brass
  - Stainless steel
15. Name the following.
- A compound that is added to lower the fusion temperature of the electrolytic bath in the extraction of aluminium.
  - The process of heating an ore to a high temperature in the presence of air.
  - The compound formed by the reaction between calcium oxide and silica.
16. From the list of characteristics given below, select the five that are relevant to non-metals and their compounds:
- Ductile
  - Conduct electricity
  - Brittle
  - Acidic oxides
  - Basic oxides
  - Discharged at anode
  - Discharged at cathode
  - Ionic chlorides
  - Covalent chlorides
  - Reaction with dilute sulphuric acid that yields hydrogen
  - 1, 2 or 3 valence electrons
  - 5, 6 or 7 valence electrons
17. Why is sodium kept under kerosene oil?
18. Name two important methods for preventing the corrosion of iron.
19. Answer the following [ICSE 2012]
- Name the chief ore of aluminium.
  - Name the process used to concentrate the above mentioned ore.
  - Why is alumina added to cryolite in the electrolytic reduction of aluminium?
  - Give the cathode and anode reactions involved in extraction of aluminium from its above mentioned ore.
  - Name the process used for the concentration of zinc blende.
  - Name two ores from which Al is extracted by electrolysis.
20. Give the chemical name and formula of cryolite, bauxite, zinc blende, calamine and haematite.
21. What happens to Al<sub>2</sub>O<sub>3</sub> and Fe<sub>2</sub>O<sub>3</sub> when bauxite is heated with a NaOH solution?
22. Name the process used for the purification of bauxite.
23. Name the process by which Al is obtained from purified bauxite.
24. Answer the following :
- Which alloy of Al is used in the construction of the body of an aircraft?
  - In construction work, why is duralumin used rather than pure Al?
  - What do you mean by the term galvanization?
  - Mention a large-scale use of Zn.
  - With Cu, Zn forms a very important alloy. Name it.
  - Zn forms an important alloy with Cu and Sn. Name it.
  - Which metal(s) is or are added to iron to make stainless steel?

8. State the property of the metal being used in the following processes:

Use of metal	Property
Aluminium in termite welding	
Zinc in galvanization	

### G. SHORT ANSWERS

- What are metals? Give some examples.
- Find the odd one out and explain your choice (note that valency is not a criterion). Copper, lead, zinc, mercury [ICSE 2009]
- What are non-metals? Give five examples.
- Write the purpose of coke and lime stone.
- Name the two substances that separate at the bottom of the blast furnace.
- Give the formula for the cryolite.
- What is rust known as chemically?
- Write the formula for zinc blende.
- Write a fully balanced equation for the reaction that takes place when magnesium is treated with dilute hydrochloric acid. [ICSE 2009]
- Compare the properties of a typical metal and a non-metal on the following criteria.
  - Electronic configuration
  - Nature of oxide
  - Oxidising or reducing behaviour
  - Conduction of heat and electricity
- What is the purpose of calcining an ore?
- What is the purpose of roasting an ore?
- What is the use of duralumin and its composition?
- Complete the following
 
$$\text{O}_2 + \text{coke} \rightarrow \text{gas Y}$$

$$\text{Gas Y} + \text{coke} \rightarrow \text{gas Z}$$

$$\text{Gas Z} + \text{iron oxide} \rightarrow \text{gas Y} + \text{Iron}$$
- Compare the properties such as electronic configuration, nature of oxides of metals and non-metals.

- What is the chemical composition of rust?
- What is meant by slag? Write its uses.
- Write the reaction between sulphur and iron filings.
- Write the equation for the action of heat on  $\text{Al(OH)}_3$ .
- How are impurities removed from bauxite?
- What is the most important chemical step in the extraction of a metal? Show how this step is carried out in the extraction of Al.
- What is the function of cryolite in the extraction of Al, other than dissolving bauxite?
- Explain the significance of each substance listed below in the extraction of Al. (a) Bauxite (b) NaOH (c) Cryolite (d) Graphite
- What is pig iron and how is it converted into steel?
- What is meant by tempering of steel?
- In order to obtain 1 ton of Al by the electrolytic process, 4 tons of bauxite, 150 kg of NaOH and 600 kg of graphite are required. Explain why so much graphite is required for the process.
- Give the equation for the reaction taking place at the cathode during the electrolytic extraction of Al.
- What is the difference between calcination and roasting ?

### H. LONG QUESTIONS

- Compare any five physical characteristics of metals with those of non-metals.
- Describe any two methods of enrichment of ores.
- Draw a neat and well-labeled diagram for the silver plating on an iron spoon. [ICSE 2012]
- A metal article is to be electroplated with silver. The electrolyte selected is sodium argentocyanide.

[ICSE 2009]

	Electrolyte	Name of cathode	Name of anode	Product at cathode	Product at anode
1.	$\text{CuSO}_4$ (aq)	Copper	Copper		
2.	$\text{PbBr}_2$ (molten)	Platinum	Platinum		

- What kind of salt is sodium argentocyanide?
- Why is it preferred to silver nitrate as an electrolyte?

(iii) State one condition to ensure that the deposit is smooth, firm and long lasting.

(iv) Write the reaction taking place at the cathode.  
(v) Write the reaction taking place at the anode.

### RECENT BOARD QUESTIONS

#### Question 1.

(a) Choose the correct answer from the options given below :

- (i) The metals zinc and tin are present in the alloy :[2013]
  - (A) Solder
  - (B) Brass
  - (C) Bronze
  - (D) Duralumin.
- (ii) When fused lead bromide is electrolyzed we observe : [2014]
  - (A) a silver grey deposit at anode and a reddish brown deposit at cathode.
  - (B) a silver grey deposit at cathode and a reddish brown deposit at anode.
  - (C) a silver grey deposit at cathode and reddish brown fumes at anode.
  - (D) silver grey fumes at anode and reddish brown fumes at cathode.
- (iii) The main ore used for the extraction of iron is : [2014]
  - (A) Haematite
  - (B) Calamine
  - (C) Bauxite
  - (D) Cryolite

(iv) The electrolyte used for electroplating an article with silver is : [2014]

- (A) silver nitrate solution
- (B) silver cyanide solution
- (C) sodium argentocyanide solution
- (D) nickel sulphate solution.

(v) Aluminium powder is used in thermite welding because : [2014]

- (A) it is a strong reducing agent.
- (B) it is a strong oxidising agent.
- (C) it is corrosion resistant.
- (D) it is a good conductor of heat.

(vi) This is not an alloy of copper : [2014]

- (A) Brass
- (B) Bronze
- (C) Solder
- (D) Duralumin

(vii) The two main metals in bronze are : [2016]

- (A) Copper and zinc
- (B) Copper and lead
- (C) Copper and nickel
- (D) Copper and tin

(b) Answer the following questions : [2011,12,14,16]

- (i) Name a metal which is found abundantly in the earth's crust.
- (ii) What is the difference between calcination and roasting ?
- (iii) Name the process used for the enrichment of sulphide ore.

- (iv) Write the chemical formulae of one main ore of iron and aluminium.
- (v) Write the constituents of electrolyte for the extraction of aluminium.
- (vi) A black powdery substance used for the reduction of zinc oxide during its extraction.
- (vii) The substance is an alloy of zinc, copper and tin.
- (viii) A metal present in cryolite other than sodium.
- (ix) A metal which is unaffected by dilute or concentrated acids.
- (x) A metal present in period 3, group 1 of the periodic table.
- (xi) Electrolytic deposition of a superior metal on a baser metal.
- (xii) The method used to separate ore from gangue by preferential wetting.
- (xiii) The catalyst used in the conversion of ethyne to ethane.

(c) Match the properties and uses of alloys in List I with the appropriate answer from List 2.

	List 1		List 2
(i)	<i>The alloy contains Cu and Zn, is hard, silvery and is used in decorative articles.</i>	A.	<i>Duralumin</i>
(ii)	<i>It is stronger than Aluminium, light and is used in making light tools.</i>	B.	<i>Brass</i>
(iii)	<i>It is lustrous, hard, corrosion resistant and used in surgical instruments.</i>	C.	<i>Bronze</i>
(iv)	<i>Tin lowers the melting point of the alloy and is used for soldering purpose.</i>	D.	<i>Stainless steel</i>
(v)	<i>The alloy is hard, brittle, takes up polish and is used for making statues.</i>	E.	<i>The metal that forms two types of ions</i>
(vi)	<i>Iron</i>	F.	<i>Solder</i>

#### Question 2.

(a) The following questions are relevant to the extraction of Aluminium : [2012,13]

- (i) State the reason for addition of caustic alkali to bauxite ore during purification of bauxite.
- (ii) Give a balanced chemical equation for the above reaction.

- (iii) Along with cryolite and alumina, another substance is added to the electrolyte mixture. Name the substance and give one reason for the addition.
- (iv) Name the other aluminium containing compound added to alumina and state its significance.
- (v) Give the equation for the reaction that takes place at the cathode.
- (vi) Explain why it is necessary to renew the anode periodically.

**Question 3.**

- (a) State your observation in each of the following cases :** [2014]

- (i) When dilute hydrochloric acid is added to sodium carbonate crystals.
- (ii) When excess sodium hydroxide is added to calcium nitrate solution.
- (iii) At the cathode when acidified aqueous copper sulphate solution is electrolyzed with copper electrodes.
- (iv) When calcium hydroxide is heated with ammonium chloride crystals.
- (v) When moist starch iodide paper is introduced into chlorine gas.

- (vi) Electricity is passed through molten lead bromide. [2013]
- (vii) Water is added to the product formed, when aluminium is burnt in a jar of nitrogen gas.

- (b) State the main components of the following alloys :** [2014]

- (i) Brass. (ii) Duralumin. (iii) Bronze

- (c) Name the following :** [2014]

- (i) The property possessed by metals by which they can be beaten into sheets.
- (ii) A compound added to lower the fusion temperature of electrolytic bath in the extraction of aluminium.
- (iii) The ore of zinc containing its sulphide.

- (d) Element X is a metal with a valency 2, Y is a non-metal with a valency 3.** [2015]

- (i) Write an equation to show how Y forms an ion.
- (ii) If Y is a diatomic gas, write an equation for the direct combination of X and Y to form a compound.
- (iii) When barium chloride solution is added to salt solution E a white precipitate insoluble in dilute hydrochloric acid is obtained. Identify the anion present in it.



# 8

# STUDY OF COMPOUNDS

## LEARNING OUTCOMES

- 8.0. Hydrogen Chloride
- 8.1. General Preparation of Hydrogen Chloride Gas
- 8.2. Physical Properties
- 8.3. Chemical Properties of Hydrogen Chloride Gas
- 8.4. Hydrochloric Acid
- 8.5. Properties of Hydrochloric Acid
- 8.6. Uses of Hydrochloric Acid
- 8.7. Ammonia
- 8.8. Laboratory Preparation of Ammonia
- 8.9. Manufacture of Ammonia (Haber Process)

- 8.10. Physical Properties of Ammonia
- 8.11. Chemical Properties of Ammonia
- 8.12. Uses of Ammonia
- 8.13. Manufacture of Nitric Acid
- 8.14. Nitric Acid
- 8.15. Properties of Nitric acid Physical Properties
- 8.16. Aqua Regia
- 8.17. Uses of Nitric Acid
- 8.18. Sulphuric Acid
- 8.19. Large Scale Preparation
- 8.20. Properties
- 8.21. Uses of Sulphuric Acid

## 8.0. HYDROGEN CHLORIDE

### Introduction

Chlorine is an important element in the halogen family. All the halogen atoms have 7 electrons in their outermost shell. The chemical reactivity of halogens decreases from fluorine to iodine. During chemical reactions, each of them gains an electron and forms an ion with a charge of -1. Similarly, the hydrogen combines with chlorine to form hydrogen chloride.

Molecular formula : HCl

Molecular mass : 36.5

Bond : covalent

*Hydrochloric acid was first discovered around 800 C.E. by the alchemist Jābir ibn Ḥayyān (Geber), by mixing common salt with vitriol (sulphuric acid).*

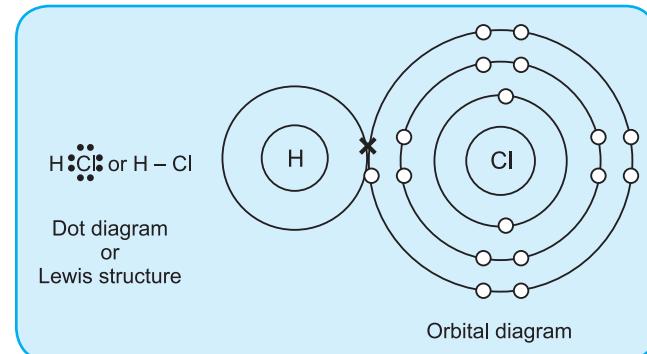


Fig. 8.1 : HCl Electron Dot Diagram and Orbital Diagram

### Occurrence

- Occurs in free state during volcanic emission.
- Glauber prepared the acid in 1648 by heating common salt (NaCl) with concentrated sulphuric acid.

## 8.1. GENERAL PREPARATION OF HYDROGEN CHLORIDE GAS

### By Synthesis

$H_2$  and  $Cl_2$  combines explosively on being kindled or exposed to sun light to form HCl.

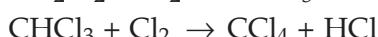
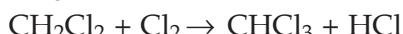
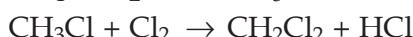
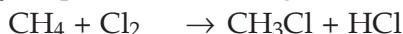


In case of activated carbon as catalyst in the dark space, the reaction occurs rapidly. This is because the activated carbon absorbs hydrogen, which increases reactivity.

Burning jet of hydrogen also burns in chlorine forming hydrogen chloride.

### As a by-product in the chlorination of hydrocarbons

HCl gas is obtained on the chlorination of hydrocarbon, in diffused sunlight e.g., methane. If you notice, one hydrogen atom of methane is replaced at every step in the following reactions.



### Laboratory Preparation Of Hydrogen Chloride Gas

#### Principle

HCl can be prepared by the action of concentrated  $H_2SO_4$  on metal chlorides.

#### Procedure

Hydrogen chloride is an important halide of hydrogen. The molecular formula is HCl and molecular weight is 36.5. It can be prepared in the laboratory by the following method : Powdered sodium chloride is taken in a round bottom flask and the apparatus is fitted as in the figure.

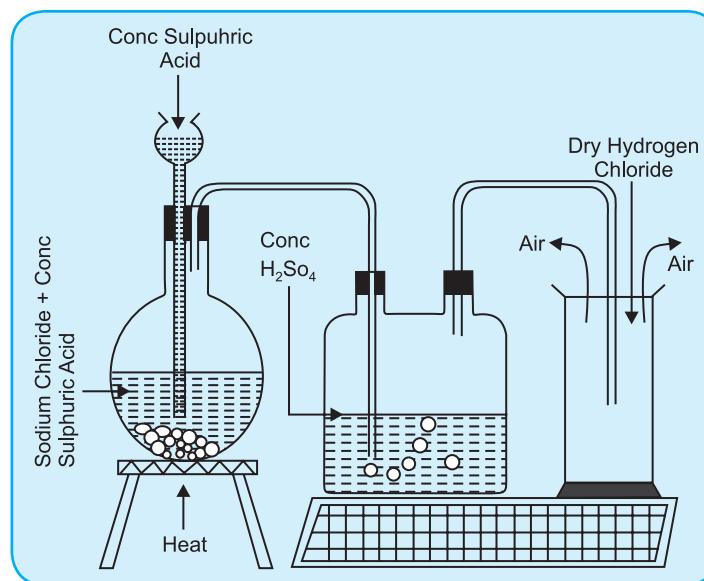


Fig. 8.2 : Preparation of HCl

Concentrated sulphuric acid is poured through the funnel present in the thistle funnel on the common salt, and they react to give off hydrogen chloride.

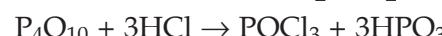
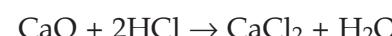
#### Reaction



#### Drying

The gas is dried by bubbling it through concentrated  $H_2SO_4$  or by passing it over anhydrous  $CaCl_2$ .

It cannot be dried using quicklime ( $CaO$ ) or  $P_4O_{10}$  as it reacts with both.



*Hydrogen is the most common material in the universe corresponding to 90% of the weight of the universe.*

#### Collection

- The gas is collected by the upward displacement of air as it is 1.28 times heavier than air.
- It cannot be collected over water, as it is highly soluble in water. An aqueous solution of hydrogen chloride is called hydrochloric acid.

#### Identification

When the jar is filled with the gas, white fumes appear. With atmospheric moisture, the gas forms a mist of small droplets of concentrated hydrochloric acid, which appear as white fumes. To check the presence of HCl, we can bring a rod dipped in ammonium hydroxide near its mouth. By doing this, we can find dense white fumes of ammonium chloride being produced.

### 8.2. PHYSICAL PROPERTIES

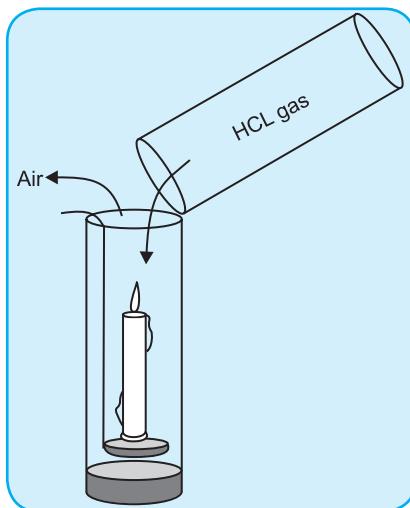
- HCl is a colourless gas with a pungent odour.
- It is neither combustible nor a supporter of combustion.
- Its density at S.T.P is  $1.6392\text{ gL}^{-1}$  and it is, therefore, heavier than air (density =  $1.29\text{ gL}^{-1}$ )
- It is not easily liquefied. The liquid boils at  $-85^\circ C$ .
- It is highly soluble in water, 1 volume of water dissolving about 460 volumes of HCl gas at  $15^\circ C$ . An aqueous solution turns blue litmus red.
- It is soluble in organic solvents such as benzene and toluene. These solutions have no effect on litmus.

#### Experiment To Demonstrate that HCl is Heavier than Air

Pour a jar of HCl gas into an 'empty' jar as if pouring water. Test the lower jar with a glass rod

moistened with liquor ammonia. White fumes are formed indicating that the lower jar contains HCl gas.

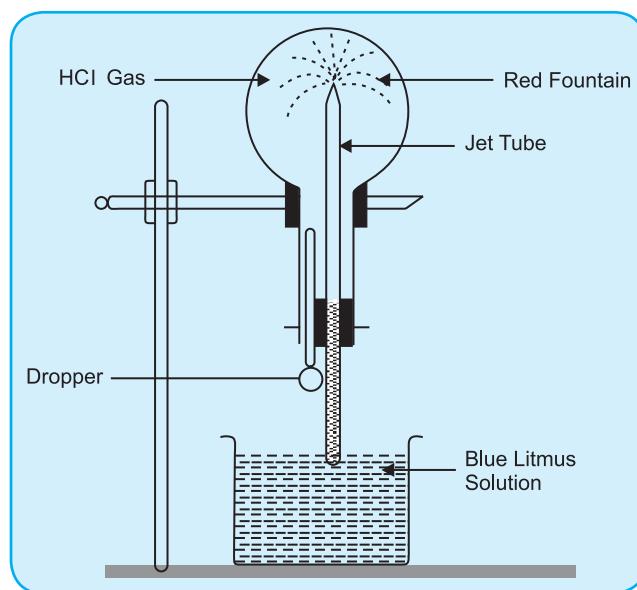
Similarly test for HCl in the upper jar. No white fumes are obtained indicating that the upper jar no longer contains HCl gas. This shows that HCl gas is heavier than air.



**Fig. 8.3 : Experiment to Determine HCl is Heavier than Air**

#### Experiment to Demonstrate the High Solubility In Water – Fountain Experiment

A round bottom flask fitted with a two-holed rubber stopper is set up as shown in the fig 8.4. A long tube, the tip of which is drawn into a jet, and an ink filler filled with water are inserted into the two holes. The other end of the long tube is immersed in a beaker of water to which a small amount of blue litmus is added. The flask is filled with dry hydrogen chloride gas. As the rubber, tube with the water is squeezed; water is forced into the flask. A red fountain is seen in the flask.



**Fig. 8.4 : Fountain Experiment**

This is because as water is squeezed into the flask, HCl gas dissolves in this water thereby reducing the pressure in the flask. The blue solution from the beaker rushes in the flask and turns red due to the acidic nature of HCl.

#### NOTE

Due to the high solubility, hydrogen chloride gas fumes in moist air forming tiny droplets of hydrochloric acid.

### 8.3. CHEMICAL PROPERTIES OF HYDROGEN CHLORIDE GAS

#### 1. Combustibility

The gas is neither combustible nor a supporter of combustion. It does not burn; rather it extinguishes a burning splint.

#### 2. Thermal Dissociation

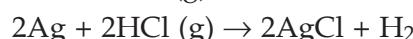
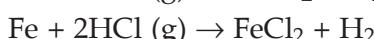
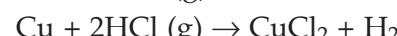
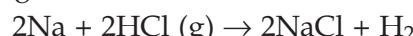
On heating above 500°C, it dissociates into hydrogen and chlorine.



#### 3. Reaction with Metals

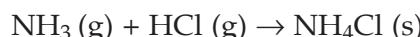
HCl gas reacts with metals only at very high temperatures, liberating hydrogen. Metals such as Cu and Ag also react under these conditions.

Burning sodium burns with a golden yellow flame in dry HCl gas.



#### 4. Reaction With Ammonia

HCl gas combines with NH<sub>3</sub> to form dense white fumes of NH<sub>4</sub>Cl, which settle in solid form.



### 8.4. HYDROCHLORIC ACID

#### Laboratory Preparation

An aqueous solution of HCl gas is called hydrochloric acid. It can be obtained by dissolving the HCl gas prepared in the laboratory in water.

The gas is passed into water until no more gas is absorbed. The product is concentrated and contains about 36% of hydrogen chloride by mass.

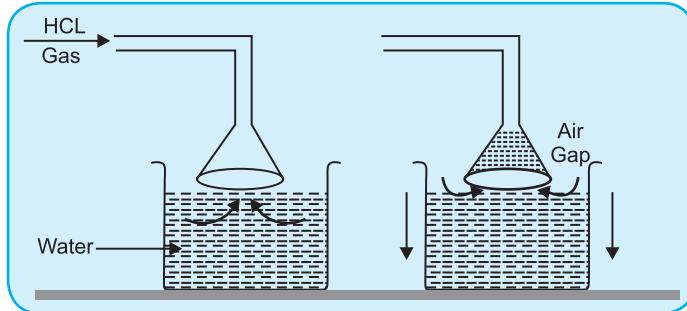
#### Procedure

An inverted funnel is used to dissolve the HCl gas

in to the water, and a trough of water is taken. During the experiment, place the funnel which passes the HCl on the surface of the water in the trough. Since, the HCl gas is highly soluble in water, the water rises up in the funnel, a reverse suction occurs and in turn, the level outside the funnel falls creating an air gap between the rim of the funnel and the surface of water.

The pressure outside and inside then becomes equal and water which had risen in the funnel as a result falls down again.

This process is done until the water in the trough is saturated with hydrogen chloride gas resulting in the formation of hydrochloric acid.



**Fig. 8.5 : Laboratory Preparation of HCl Acid**

- The funnel arrangement should.
- Prevent or minimize back suction of water.
- Provide a large surface area for absorption of HCl gas.

### Precaution

To prevent accident, an empty flask is fitted in-between the generative flask and the water trough. In case, the back suction occurs, the water will collect in it and will not reach the generating flask.

HCl acid forms constant boiling mixture at 110°C.

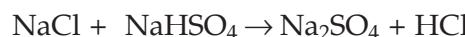
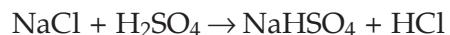
A dilute aqueous solution of hydrochloric acid gets gradually concentrated on distillation, until the concentration of the acid reaches 22.2% HCl by weight and 77.8% water by weight, which boils at 110°C.

**Azeotropic mixture** is a mixture of two liquids, which boils at a constant temperature and distils over at the same temperature without change in its composition.

### Leblanc Process

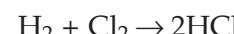
Hydrochloric acid is obtained by the action of hot concentrated sulphuric acid on common salt. Both these substances are taken in an iron pan and heated. At the initial stage HCl and NaHSO<sub>4</sub> are obtained, which (NaCl + NaHSO<sub>4</sub>) is heated more strongly. HCl and Na<sub>2</sub>SO<sub>4</sub> (salt cake) are formed. The HCl gas formed in both the stages is led into a brickwork or

stone tower packed with coke. The gas is absorbed by the water trickling from the top of the tower to form a solution of hydrochloric acid.



### Synthesis From H<sub>2</sub> And Cl<sub>2</sub>

H<sub>2</sub> and Cl<sub>2</sub> are obtained at the cathode and anode, respectively, in the electrolysis of brine (NaCl solution) during the manufacture of caustic soda (NaOH). HCl gas is made by burning a mixture of H<sub>2</sub> and Cl<sub>2</sub> and dissolved in water to obtain the acid.



## 8.5. PROPERTIES OF HYDROCHLORIC ACID

### Physical Properties

- The pure acid solution is colourless but the commercial acid is often yellow due to the presence of FeCl<sub>3</sub> as impurity.
- It gives a pungent choking smell.
- It is sour in taste.
- It is corrosive in nature.
- It is readily soluble in water in all proportions.
- The boiling point of the acid varies with its concentration. When the acid content is 20.24% by mass, the boiling point is constant at 110°C.

### Chemical Properties

#### 1. Nature

Aqueous solution is strongly acidic, and shows all properties of acids.

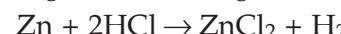
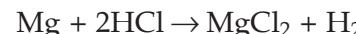
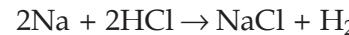
Action on indicators

**Table 8.1 Action on indicators**

Indicator	Original colour	Changes
Moist litmus	Blue	Red
Methyl orange	Orange	Pink
Phenolphthalein	Colourless	Colourless

#### 2. Action on Metal

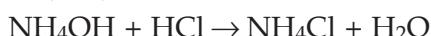
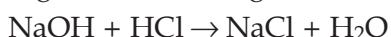
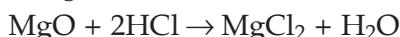
Concentrated hydrochloric acid dissolves many metals, generally those elements, which are situated above hydrogen in the electrochemical series forming metallic chlorides and liberating hydrogen gas.



#### 3. Action on Oxides and Strong Hydroxides

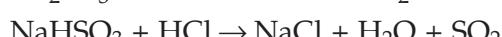
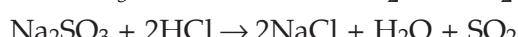
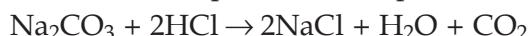
Hydrochloric acid is a strong acid, obtained by dissolving hydrogen chloride gas in water. This

solution contains  $\text{H}^+$  and  $\text{Cl}^-$  ions and can conduct electricity. It is a strong acid, reacting with oxides and strong alkalis to give salts and water.



#### 4. Action on Salts of Weaker Acids :

It decomposes salts of weaker acids, e.g., carbonates, hydrogen carbonates, sulphites and sulphides.



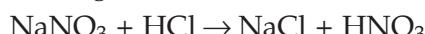
#### 5. Action on Thiosulphates :

Dilute HCl reacts with thiosulphates to give  $\text{SO}_2$  gas on warming and the solution turns turbid due to the precipitation of sulphur.



#### 6. Reaction with Nitrates :

Concentrated HCl gives reddish brown vapours of  $\text{NO}_2$  on heating with nitrates.



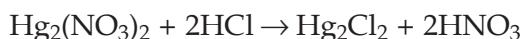
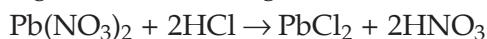
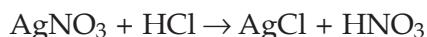
#### 7. Reaction with Strong Oxidising Agent

In the reaction with strong oxidising agents, it functions as a reducing agent, forming  $\text{Cl}_2$ .



#### 8. Precipitation of Some Metal Chlorides :

$\text{AgCl}$ ,  $\text{PbCl}_2$  and  $\text{Hg}_2\text{Cl}_2$ , being insoluble, are precipitated when hydrochloric acid is added to solutions containing  $\text{Ag}^+$ ,  $\text{Pb}^{2+}$  or  $\text{Hg}^{2+}$  ions.  $\text{PbCl}_2$  dissolves in hot water and on cooling with tap water it reappears.



#### NOTE

When silver chloride is exposed to light, it decomposes into chlorine and metallic silver, which is liberated as a fine black powder.

#### 9. Formation of Aqua Regia

A mixture of hydrochloric acid and nitric acid (3:1 by volume) is known as aqua regia.



The liberated nascent chlorine reacts with the noble metals like gold and platinum to give their soluble chlorides.

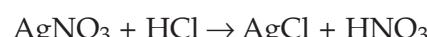


#### 8.6. USES OF HYDROCHLORIC ACID

- It is used as a laboratory reagent and for the preparation of aqua regia mixture.
- For pickling of rust on steel before the galvanization process.
- For cleaning the metal surfaces before painting, electroplating, soldering etc. For decomposing bones to make gelatine.
- For refining oils, fats and waxes.
- In the tanning of leather.
- In the manufacture of dyes and textiles.
- Dilute HCl is prescribed to patients with decreased activity of their gastric juices.

#### Tests For Hydrogen Chloride And Hydrochloric Acid

- HCl gas gives a characteristic irritating smell.
- A glass rod dipped in concentrated HCl when brought close to ammonia, causes dense white fumes.  $\text{HCl}(\text{g}) + \text{NH}_3(\text{g}) + \text{NH}_4\text{Cl}(\text{s})$
- When a little amount of silver nitrate solution is treated with HCl acid as well as HCl gas, both give a white precipitate of silver chloride, which is insoluble in  $\text{HNO}_3$  but soluble in  $\text{NH}_3$  solution.



- When concentrated hydrochloric acid is treated with the oxidising agents like manganese dioxide, a greenish yellow gas is evolved which turns starch iodide paper blue black.





## 8.7. AMMONIA

### Introduction

Nitrogen is a gas. It is a diatomic element, and is present in the free state in the air along with oxygen, carbon dioxide, and the rare gases. It is the one of the principal elements found in all living organisms. Some of the nitrogen compounds like ammonia ( $\text{NH}_3$ ) and the oxides, though very useful, are atmospheric pollutants too.

Let us study the chemistry of two important compounds of nitrogen—ammonia and nitric acid—in this chapter.

**Ammonia has a sharp odour and is used in fertilizers, refrigeration, and cleaning products. Ammonia is a colourless gas with a sharp, pungent odour. It is both manufactured and also produced naturally by bacteria, decaying plants and animals, and animal waste.**

### Ammonia

Ammonia occurs in traces in the atmosphere. Gaseous ammonia was first prepared by Priestly. Davy established its formula ( $\text{NH}_3$ ) in 1800.

In the formation of ammonia, an atom of N combines with three atoms of H by sharing three of its valence electrons, one with each H atom.

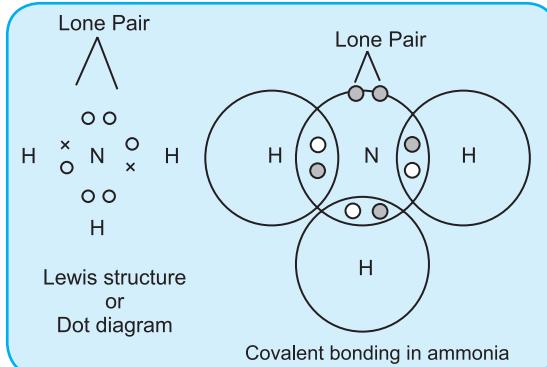


Fig. 8.6 : Ammonia Electron Dot Structure and Covalent Bonding

MOLECULAR FORMULA :  $\text{NH}_3$

RELATIVE MOLECULAR MASS : 17

### Occurrence

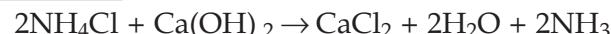
Ammonia is present in small amounts in air and in traces in natural water.

The emission of direct and indirect greenhouse gases, namely methane ( $\text{CH}_4$ ), carbon dioxide ( $\text{CO}_2$ ), ammonia ( $\text{NH}_3$ ), nitric ( $\text{NO}$ ) and nitrous oxide ( $\text{N}_2\text{O}$ ) was studied as greenhouse gas.

## 8.8. LABORATORY PREPARATION OF AMMONIA

Ammonia can be conveniently prepared in the laboratory by heating ammonium chloride with slaked lime (calcium hydroxide) in 1:3 ratio in a hard glass tube. The ammonia gas evolved is passed through a tower packed with quick lime and collected by downward displacement of air. Ammonia is highly soluble in water; hence, it cannot be collected over water, it can, however, be collected over mercury.

### Reaction



When ammonia enters the body as a result of breathing, swallowing or skin contact, it reacts with water to produce ammonium hydroxide. This chemical is very corrosive and damages cells in the body on contact. To reduce the effects from exposure to ammonia, it is important to wash eyes and skin as quickly as possible with large amounts of water.

### Precaution

The flask is tilted to avoid the breakage of heated flask because the water formed during the reaction will enter back to the reaction flask.

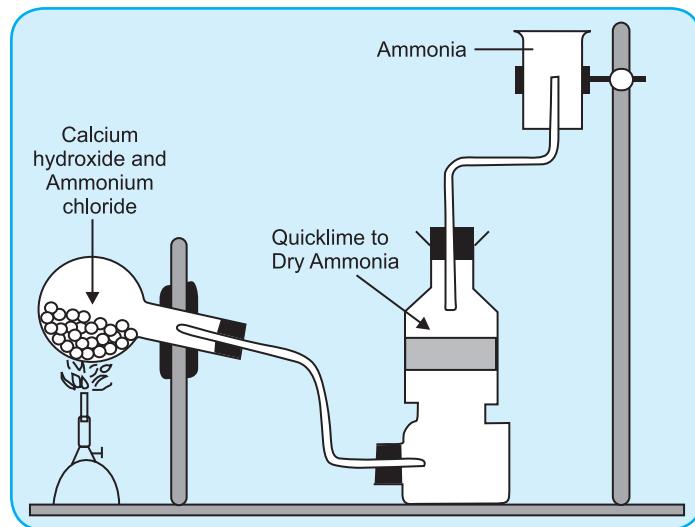


Fig. 8.7 : Preparation of Ammonia

### Drying of Ammonia Gas

Drying of ammonia gas is done by passing the gas in to a drying tower containing lumps of quick lime ( $\text{CaO}$ ).

Other drying agents like concentrated sulphuric acid, phosphorous pentoxide and anhydrous calcium

chloride are avoided. The reason is since ammonia is basic in nature it readily undergoes reaction with acids.

### NOTE

- Ratio of reactants taken is 1 : 3 by weight, higher ratio by weight of the alkali may contract the loss by sublimation of  $\text{NH}_4\text{Cl}$ .
- Calcium hydroxide is used, as it is cheap and not deliquescent like other alkalis.
- Though all ammonium salts, on heating with alkalis, give  $\text{NH}_3$ ,  $\text{NH}_4\text{NO}_3$  is not used as it is explosive in nature and it decomposes forming nitrous oxide and water vapour.

### From Metal Nitrides

Ammonia can also be obtained by the action of warm water on nitrides of metals like magnesium or aluminium.

### Procedure

Magnesium nitride or aluminium nitride is kept in a round bottomed flask and hot water is added gradually through a thistle funnel and a delivery tube as shown in the fig 8.11.

Between both the ammonia gas and magnesium hydroxide formed in the flask only the ammonia gas is liberated and escapes through the delivery tube.

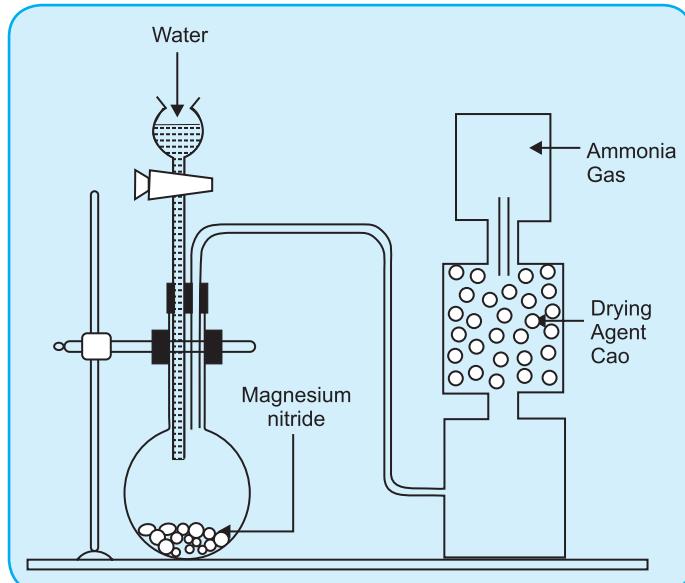


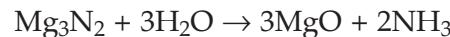
Fig. 8.8 : Preparation of Ammonia from Magnesium Nitride

The ammonia gas formed is collected by holding the jar with its mouth downwards over the delivery

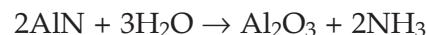
tube. The gas can be dried by passing it through quicklime packed in a drying tower.

### Reaction

$\text{Mg}_3\text{N}_2$  and the other group IIA nitrides give  $\text{NH}_3$  with hot or boiling water.



AlN gives  $\text{NH}_3$  on being heated with steam.



### Preparation of Aqueous Ammonia

Aqueous ammonia is prepared in the laboratory by dissolving the ammonia gas in water; the back suction is possible while using a funnel with a suction tube, to dissolve the gas in the water. The set up is similar to that used in the preparation of HCl gas.

### Procedure

Water is taken in a trough and the mouth of the funnel is made to be immersed in water. When the ammonia gas get dissolved in water at a higher rate than its production in the flask, the pressure in the funnel above the water level decreases for some time and water rushes into the funnel. Therefore, the rim of the funnel loses its contact with water. Since ammonia produced pushes the water down, the funnel comes in contact with water again.

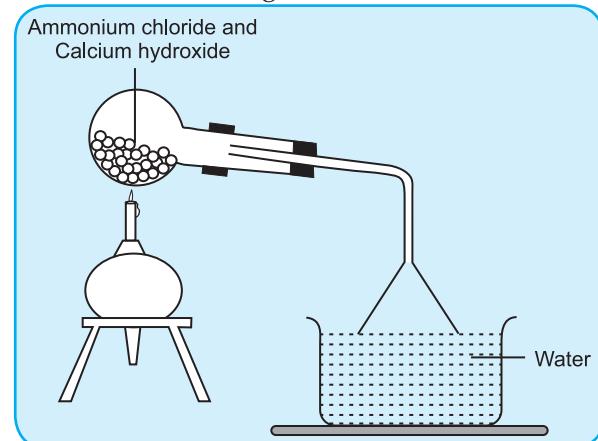


Fig. 8.9 : Preparation of Aqueous Ammonia

### 8.9 MANUFACTURE OF AMMONIA (HABER PROCESS)

### Principle

The Haber process manufactures a large scale of ammonia from  $\text{N}_2$  and  $\text{H}_2$ .

$\text{NH}_3$  is formed on heating a mixture of  $\text{N}_2$  and  $\text{H}_2$ . The reaction is reversible and exothermic.



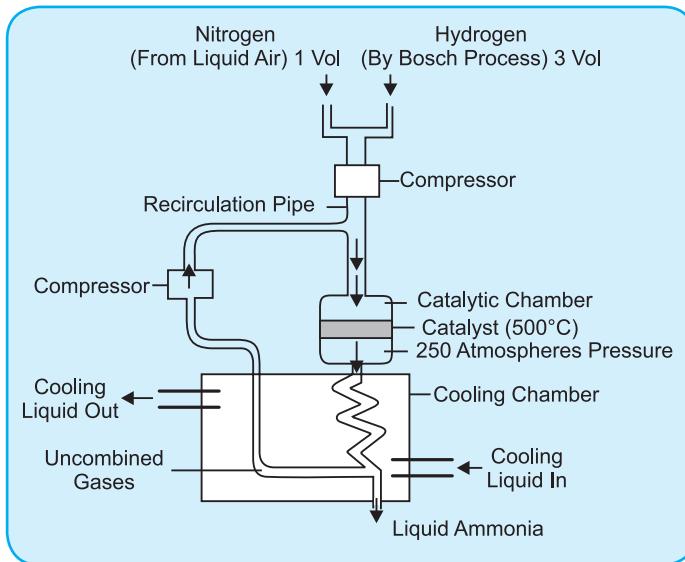


Fig. 8.10 : Haber Process

### Favourable Condition

- A mixture containing N<sub>2</sub> and H<sub>2</sub> in the volume ratio 1 : 3 is taken.
- The temperature 720 - 770 K is to be maintained.
- Finely divided iron is used as a catalyst and molybdenum as promoter.
- The ammonia formed is liquefied and removed from the equilibrium mixture.

**According to Le Chatelier's principle favourable conditions for maximum yield are :**

### Low Temperature :

The optimum temperature for the reaction is 720–770 K. The low temperature favours the formation of ammonia on a large scale during the process because, in the Haber process the formation of NH<sub>3</sub> forward reaction, the heat is liberated. This leads to increase in temperature, but in the case of backward reaction,

i.e., decomposition of NH<sub>3</sub>, the fall in temperature of the equilibrium mixture takes place. So, the higher the temperature, lower the yield. Conversely, lower the temperature higher the yield of ammonia. However, at low temperature the reaction is slow.

### High Pressure :

According to the kinetic theory of gases, at constant temperature, the pressure of the gas is directly proportional to the number of molecules striking the wall of the container. When a high pressure in the order of 200–900 atmospheres is applied, more amount of ammonia is formed, because the number of molecules decrease when ammonia is formed, and so does the pressure.

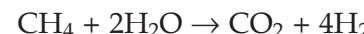
### Increase in the Reaction Velocity :

A catalyst changes the rate of the reaction without itself undergoing a net chemical change. During a reaction, a catalyst changes the rates of the forward reaction and the backward reaction at the same time. Therefore, it helps to attain equilibrium within a shorter time. A finely divided iron with molybdenum or calcium acts as promoter.

Since the reaction is exothermic, the heat evolved maintains the temperature. External heating is required after the reactants are initially heated.

### Process

N<sub>2</sub> and H<sub>2</sub> used in the Haber process must be purified in dry condition to avoid the catalyst to be poisoned. H<sub>2</sub> is produced on a large scale by passing a mixture of methane and steam over a nickel catalyst at 750°C.



A calculated amount of air is mixed with the product so that the O<sub>2</sub> of the air reacts with a part of the H<sub>2</sub> and leaves N<sub>2</sub>. A mixture containing N<sub>2</sub> and H<sub>2</sub> in the volume ratio 1:3 is then formed. The CO<sub>2</sub> present in the mixture is removed by dissolution in a concentrated solution of K<sub>2</sub>CO<sub>3</sub>.

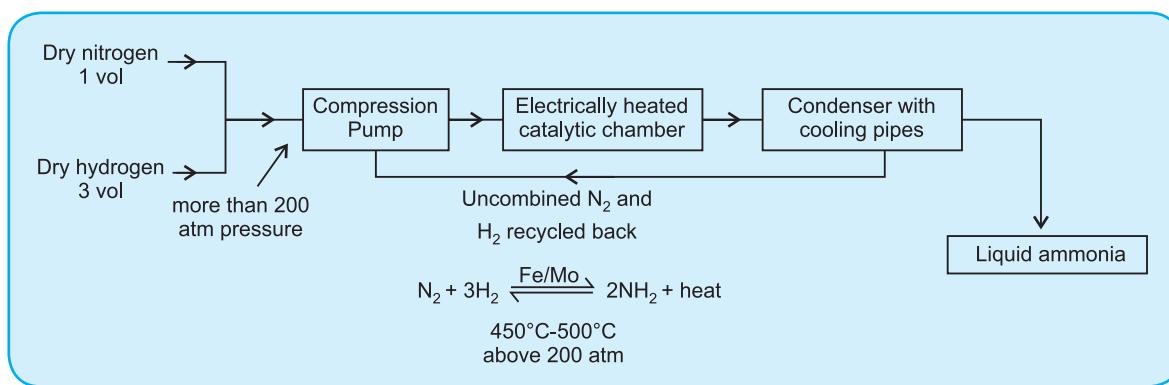


Fig. 8.11 : Flow Chart of Haber Process

A mixture of pure N<sub>2</sub> and H<sub>2</sub> in the volume ratio 1 : 3 is passed under a pressure of 200–900 atm

through a reaction chamber containing finely divided iron as catalyst with molybdenum as promoter. The

equilibrium mixture is cooled to liquefy  $\text{NH}_3$ , and the residual mixture of  $\text{N}_2$  and  $\text{H}_2$  gases is recycled.

### 8.10 PHYSICAL PROPERTIES OF AMMONIA

- (a) Ammonia is a colourless pungent-smelling gas. Vapours of ammonia evolving from concentrated ammonia solutions are severely irritating to the eyes and the respiratory tract. If inhaled for a longer time, it may cause death because it is a heart stimulant; hence, these solutions should be handled with care, and if possible, placed inside a fume cupboard.
- (b) It is less dense (vapour density = 8.5) than air and is collected by the downward displacement of air.
- (c) It is alkaline in nature.
- (d) It has a boiling point of  $-33.5^\circ\text{C}$  and freezing point is  $-77.7^\circ\text{C}$ .
- (e) It is extremely soluble in water. One volume of water dissolves about 739 parts of the gas at  $20^\circ\text{C}$ . Its high solubility is demonstrated by the fountain experiment.
- (f) It can be easily liquefied.

#### The Fountain Experiment

Collect ammonia gas in a clean dry round-bottom flask. Close it with a two-holed rubber stopper and introduce a straight glass tube. Invert the apparatus and fix it to a stand. Introduce the other end of the delivery tube into beaker containing water. Now through a dropper introduce a few drops of phenolphthalein into the flask and close the hole.

What do you observe ?

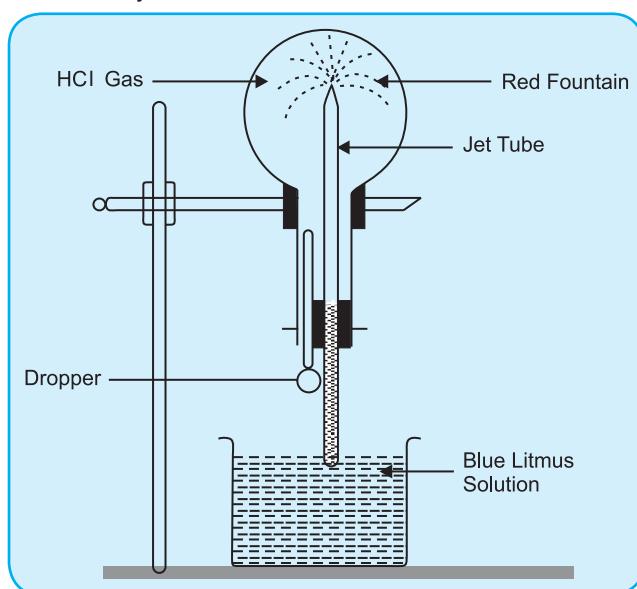


Fig. 8.12 : Fountain Experiment

You will observe that the ammonia in the flask dissolves in water. This reduces the pressure in the flask. Subsequently, phenolphthalein is forced into the flask due to pressure and flows as a pink fountain. This experiment shows the high solubility of ammonia in water.

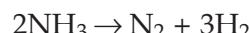
#### NOTE

A bottle of liquor ammonia should be opened very carefully, only after cooling it in ice or cold water. There is high pressure inside the bottle of ammonia and on cooling, the pressure decreases. This prevents sudden flushing out of the gas.

### 8.11. CHEMICAL PROPERTIES OF AMMONIA

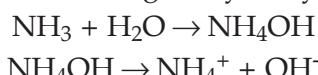
#### 1. Dissociation

Ammonia dissociates into nitrogen and hydrogen at red heat or when an electric discharge is passed through it.



#### 2. Basic Properties

Due to covalent nature, ammonia is neutral even in liquid form. In addition, it is weak base; the basicity is due to a lone pair of electrons on its nitrogen atom. It partially dissociates to give hydroxyl ions.



#### Action on Indicators

Since aqueous ammonia solution is a base, it turns red litmus blue and Methyl orange to yellow

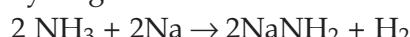
#### 3. Reaction with Acids

Since ammonia is a base, it combines with acids to form salts.

Ammonia gas + acid $\rightarrow$ ammonia salt		
$\text{NH}_3$	$+$	$\text{HCl} \rightarrow \text{NH}_4\text{Cl}$
$\text{NH}_3$	$+$	$\text{HNO}_3 \rightarrow \text{NH}_4\text{NO}_3$
$2\text{NH}_3$	$+$	$\text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4$

#### 4. Reaction with Metals

When ammonia is passed over heated sodium or potassium, the corresponding metal amides are produced. Hydrogen is liberated in this reaction.



#### 5. Action with Halogens :

Both chlorine and bromine oxidize ammonia to nitrogen. The corresponding halogen acids are formed.





### 6. Reaction with Oxygen :

Ammonia burns in oxygen with yellowish green flame and does not burn in air. It is not a supporter of combustion.

Set up the apparatus as shown in the figure 8.16. Two inlet tubes are enclosed in a wider tube. Pack the bottom of the wide tube loosely with cotton wool. Send in a slow stream of oxygen. The cotton wool helps the oxygen to spread. Also, send a stream of NH<sub>3</sub> by heating an ammonia solution separately. After a minute, when the ammonia begins to smell, bring a lighted taper to the mouth of the ammonia tube. Ammonia burns with a small yellowish brown flame and produces water vapour and nitrogen.

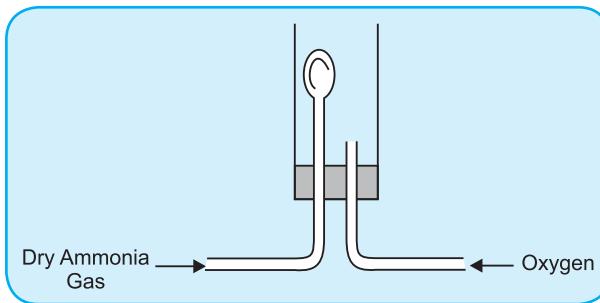
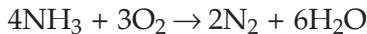


Fig. 8.13 : Mixture of Ammonia Gas and Oxygen



The above reaction is exothermic and irreversible.

The mixture of ammonia and oxygen are explosive and therefore dangerous.

### 7. The Catalytic Oxidation of Ammonia

A dry ammonia gas is passed over the hot platinum chloride which is present in the long combustion tube. Make sure that the air inlet tube must not dip into the NH<sub>3</sub> solution; otherwise too much NH<sub>3</sub> will be carried over. Heat the combustion tube until the platinum is red hot, and sucks the mixture of NH<sub>3</sub> and air into the combustion tube by turning on the suction pump. In a few seconds, brown fumes will appear in the reviver.



### 8. Reaction with Carbon Dioxide

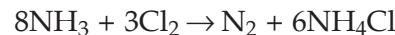
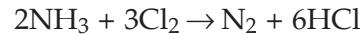
Ammonia reacts with carbon dioxide at 185°C and 200 atm to form urea (NH<sub>2</sub>CONH<sub>2</sub>), which is widely used as fertilizer.



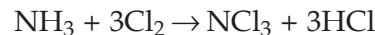
### 9. Ammonia as a Reducing Agent

- Reduction of chlorine to hydrogen chloride

When chlorine gas is passed into an excess of a concentrated NH<sub>3</sub> solution, the latter is oxidised to nitrogen.

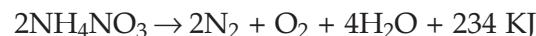


When Cl<sub>2</sub> is in excess, the explosive liquid NCl<sub>3</sub> is formed.



### 8.12. USES OF AMMONIA

- It is used as a refrigerant in ice-plants.
- It was used as a refrigerant. Evaporation of a liquid needs heat energy. About 17g of liquid NH<sub>3</sub> absorbs 5700 cal of heat from the surrounding water. This cools the water and ultimately freezes it to ice. However, its use as a refrigerant was discontinued because the leakage of NH<sub>3</sub> causes air and water pollution.
- As a refrigerant, ammonia was first replaced by chlorofluorocarbons (CFCs), e.g., Freon (CCl<sub>2</sub>, F<sub>2</sub>). However, the leakage of CFCs causes depletion of the ozone layer, which protects us from UV rays. It has been found that the chlorine present in CFCs is the main source. It is released as what is known as a free radical, reacting with ozone. So, CFCs have now been replaced by substances containing lesser or no chlorine – they are hydro chlorofluorocarbons (HCFCs, e.g., CHClF<sub>2</sub>) and hydro fluorocarbons (HFCs, e.g., CHF<sub>3</sub>).
- It is used in the manufacture of nitric acid by Ostwald's process.
- It is used in the manufacture of urea.
- Since hydrogen is highly combustible, it is transported as liquefied ammonia and then catalytically treated to obtain hydrogen.
- It is used in the manufacture of fertilizers such as ammonium phosphate, ammonium sulphate, ammonium nitrate etc. It is used in the manufacture of sodium bicarbonate and washing soda by Solvay's process.
- It is used as a cleansing agent for removal of grease in dry cleaning and cleaning liquids, etc. Ammonium nitrate, the salt formed by ammonia and nitric acid, is an explosive, decomposing above 250°C to form N<sub>2</sub>, O<sub>2</sub> and steam.



### Tests for Ammonia

- Ammonia has a distinct pungent odour.
- It turns moist red litmus blue.
- When a glass rod dipped in concentrated hydrochloric acid is brought in contact with



ammonia, thick white fumes of ammonium chloride are given out.

- When a little amount of ammonia is added to a copper sulphate solution the solution turns into blue precipitate. On addition of excess amount of ammonia, the precipitate dissolves and the solution turns into deep blue colour.
- Ammonia or ammonium ions produce a brown precipitate or brown or yellow colouration with Nessler's reagent (Nessler's reagent is a solution of potassium mercuric-iodide in potassium hydroxide).

### 8.13. MANUFACTURE OF NITRIC ACID

A German chemist Ostwald, in 1914, developed Ostwald process.

Nitric acid is mostly manufactured from ammonia by the Ostwald process. In this process, ammonia (10:1 by volume) is catalytically oxidized to nitric oxide, which is converted to nitric acid.

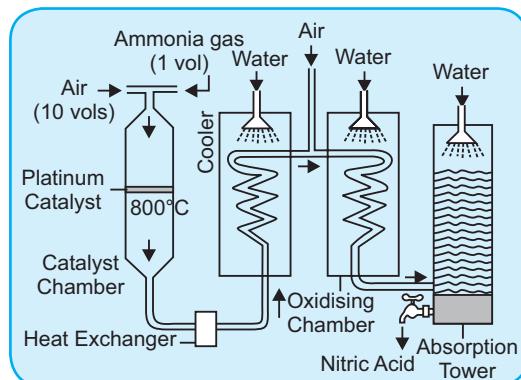


Fig. 8.14 : Haber Process

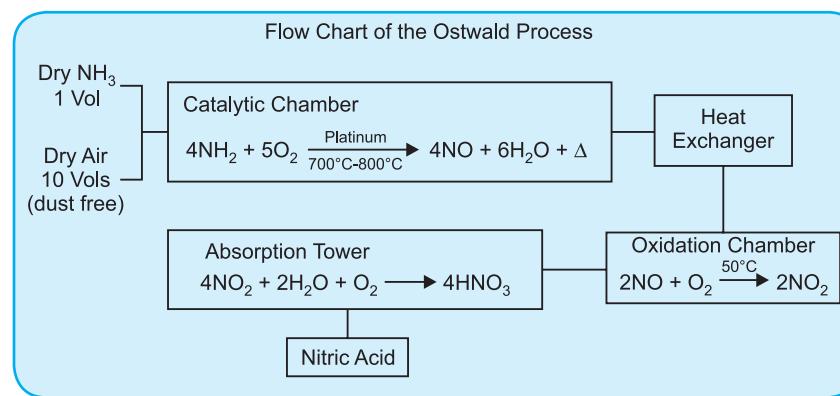


Fig. 8.15 : Flow Chart for the Contact Process

### 8.14. NITRIC ACID

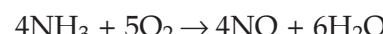
Molecular formula : HNO<sub>3</sub>

Molecular mass : 63

#### Introduction

Nitric acid is a mineral acid like sulphuric acid and hydrochloric acid. The molecular formula of nitric acid is HNO<sub>3</sub>. The mineral acids are so named because they can be obtained from naturally occurring

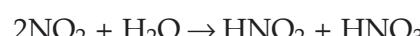
The ammonia required for the purpose is mainly obtained by Haber's process. When a mixture of ammonia and excess air is first compressed and passed into the catalytic chamber containing red-hot platinum, which is about 800°C, nitric oxide is formed.



The hot gas is passed through a heat exchanger, where they are mixed with air and passed into another cooling chamber called oxidising chamber. Then the nitric oxide, thus produced, is allowed to react with excess air to form nitrogen dioxide.



Nitrogen dioxide is then absorbed in water to get nitric acid.



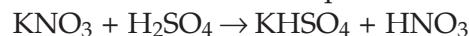
The nitric acid obtained at the bottom is concentrated above 50% and it is further distilled to get 68% nitric acid.

The fuming nitric acid (98%) is obtained by distilling the above 68% nitric acid over concentrated sulphuric acid.

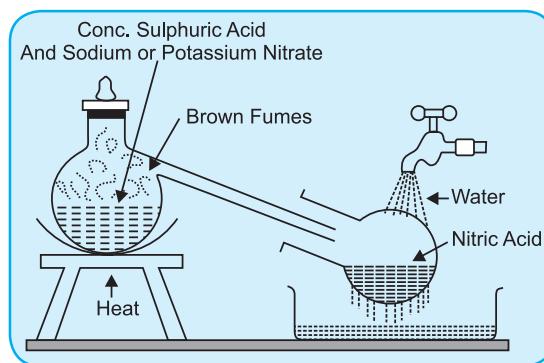
salts (minerals). In addition, the chemist gave the name *aqua fortis* to nitric acid meaning 'strong water'. Traces of nitric acid occur in atmospheric air, where it is formed by electric discharges and washed down by rain.

#### Laboratory Preparation of Nitric Acid

The nitric acid is prepared in the laboratory by the action of concentrated H<sub>2</sub>SO<sub>4</sub> on potassium nitrate.

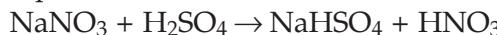


Equal parts of  $\text{KNO}_3$  and  $\text{H}_2\text{SO}_4$  are taken in terms of mass in a stopper retort. The stem of the retort is introduced into a small flask cooled under tap water. The mixture is heated gently to  $180^\circ\text{C}$ - $200^\circ\text{C}$  in a glass retort. Sulphuric acid is a non-volatile acid and produces volatile nitric acid on reacting with potassium nitrate; the nitric acid produced distils over and is collected in the flask.



**Fig. 8.16 : Preparation of Nitric Acid**

Instead of potassium nitrate, sodium nitrate ( $\text{NaNO}_3$ ) may also be used. In this case, the reaction may be represented as :



#### NOTE

Pure nitric acid is colourless but the acid obtained in the laboratory is slightly yellow. The yellow colour is due to dissolution of reddish brown coloured nitrogen dioxide gas in the acid.

- The yellow colour due to presence of  $\text{NO}_2$  gas is taken out by the dry air or  $\text{CO}_2$  gas, which is bubbled through the yellow coloured nitric acid solution.
- The yellow colour due to presence of  $\text{NO}_2$  gas is made to dissolve in the addition of excess water.

### 8.15 PROPERTIES OF NITRIC ACID

#### Physical Properties

1. Anhydrous nitric acid is a colourless liquid (98% concentration). The commercial nitric acid (68% concentration) is yellowish brown in colour.
2. It is highly corrosive, i.e., it has a tendency to destroy objects by slow oxidation.
3. It has a suffocating smell.
4. It has a sour taste.
5. It has a boiling point of  $86^\circ\text{C}$  and melts at  $-42^\circ\text{C}$ .
6. It is soluble in water in all proportion.

#### Chemical Properties

##### 1. Stability

Nitric acid is decomposed on being heated (e.g., by dropping the concentrated acid on a hot pumice stone) or on exposure to light.



Dissolution of the  $\text{NO}_2$  in the acid gives the acid a yellow colour. Only partly filled bottles acquire this colour. Nitric acid in a completely filled bottle is colourless as there is no space for the vapours to form.

#### 2. Acidic Properties

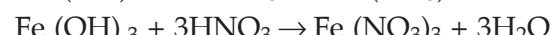
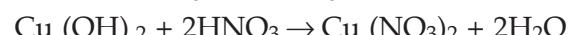
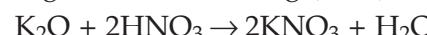
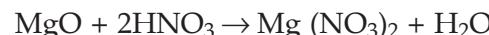
Nitric acid is a strong monobasic acid. It shows all the reactions typical of acids. It forms salts with alkalis, and reacts with carbonates and bicarbonates to form salt and water with the evolution of carbon dioxide. It attacks all the metals.

#### It Turns

- Blue litmus red,
- Methyl orange pink,
- Phenolphthalein remains colourless.

#### Reaction with Basic Metal Oxides

It reacts with basic metal oxides and hydroxides to form salts (corresponding nitrates).



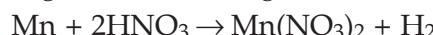
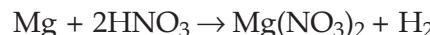
#### Reaction with Carbonates and Bicarbonates

It reacts with carbonates and bicarbonates to give salt, water and carbon dioxide



#### Reaction with Metals

On reaction with nitric acid metals usually do not give out  $\text{H}_2$ . Only Mg and Mn react with very dilute  $\text{HNO}_3$  to give  $\text{H}_2$  (1%) and the action of nitric acid on metals depends on the temperature and concentration.



It reacts with all metals except the noble metals gold and platinum.

#### 3. Oxidising Properties

Concentrated nitric acid is a powerful oxidising agent. Non-metals like sulphur, carbon and phosphorus are oxidised to their corresponding acids.



It oxidises ferrous sulphate to ferric sulphate.



Ferrous sulphate + dil. nitric acid → Ferric sulphate

### Passivity :

It can be removed, by rubbing the surface layer with sand paper, or by treating with strong reducing agents.

#### NOTE

Dilute nitric acid is generally considered a typical acid except for its reaction with metals, as it does not liberate hydrogen. It is a powerful oxidising agent and the nascent oxygen formed oxidises the hydrogen to water.

### Effects of Heat on Nitrates

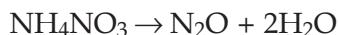
Sodium and potassium nitrates or alkali metal nitrates on heating decompose and liberate oxygen gas.



Silver nitrate and mercury nitrate decomposes to give  $\text{NO}_2$  and  $\text{O}_2$ .

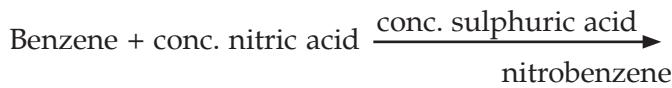


Ammonium nitrate decomposes to give water vapour and laughing gas.



### Reaction with Organic Compounds

Concentrated nitric acid in the presence of concentrated sulphuric acid, reacts with benzene to form nitrobenzene. This process is known as nitration.



### Reaction on Proteins

Nitric acid reacts with proteins to give xanthoprotein, a yellow-coloured compound. This is used as a test for proteins.

#### NOTE

Our skin turns yellow in colour when it comes into contact with concentrated nitric acid due to the formation of xanthoprotein.

### 8.16. AQUA REGIA

Another name for aqua regia is royal water. Aqua regia is a mixture of concentrated  $\text{HNO}_3$  and concentrated  $\text{HCl}$  in the ratio of 1 : 3 by volume. It is used to dissolve gold and platinum in jewellery. Royal water dissolves all metals except silver.



Here nitric acid oxidises hydrochloric acid to give their corresponding chlorine.

### 8.17. USES OF NITRIC ACID

1. In the manufacture of nitrates used as fertilizers like ammonium nitrate, calcium nitrate, etc.
2. In the manufacture of explosives like gun cotton, TNT (Trinitrotoluene), picric acid (trinitrophenol), and dynamite (trinitroglycerine).
3. It is used to make etchings on metals.
4. In the manufacture of dyes, medicines, perfumes, and artificial silk.
5. In the manufacture of sulphuric acid, by lead chamber process.
6. It is used to prepare aqua regia.
7. As an important laboratory reagent.
8. It acts as a rocket fuel oxidant.
9. In photography in the form of silver salt.

### Tests for Nitrate

#### Copper Turning Test

A mixture of copper turnings and nitrate salt is taken in a boiling test tube and heated. On heating the mixture, a brown gas is evolved. This gas is nothing but nitrogen dioxide ( $\text{NO}_2$ ). This indicates the presence of nitrate in the given salt.



#### Brown Ring Test

To a small amount of the nitrate salt solution a little amount of freshly prepared ferrous sulphate solution is added in a test tube. To this mixture, add concentrated sulphuric acid along the sides of the test tube. While cooling it under the tap a brown ring is formed at the junction of the two layers this confirms the presence of nitrate. While shaking or warming the mixture, the brown ring disappears. This test is commonly known as the brown ring test.

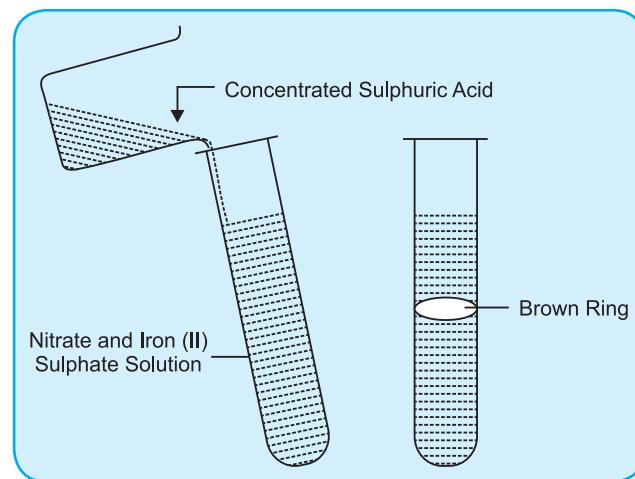
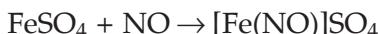


Fig. 8.17 : Brown Ring Test



The brown ring formed during the test is nitroso ferrous sulphate, which is formed in between the junction of the two liquids. Since the concentrated sulphuric acid is heavier, it settles down and ferrous sulphate being lighter floats at the top.

### 8.18. SULPHURIC ACID

Molecular formula :  $\text{H}_2\text{SO}_4$

Molecular mass : 98

#### Introduction

Sulphuric acid is probably one of the most important chemical substances not found naturally. Its manufacture is therefore important and wide. The total world production of sulphuric acid is about 25,000,000 tons a year. It is also called as the king of chemicals and is obtained in traces in the rainwater where sulphur-bearing coal is burnt. It is also known as oil of vitriol.

#### Preparation

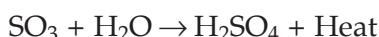
Majority of the sulphuric acid is prepared from  $\text{SO}_2$ , but increasingly, because of environmental considerations, sulphuric acid is recycled. If the levels of impurities are low, the acid is evaporated in stages to the appropriate concentration and any organic impurities are destroyed by oxidation.

#### Oxidation of S by $\text{HNO}_3$ :

Sulphur is oxidised to  $\text{H}_2\text{SO}_4$  by hot concentrated  $\text{HNO}_3$ .



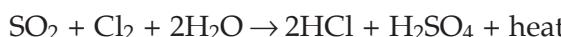
#### Oxidation of $\text{SO}_2$ :



If  $\text{SO}_2$  is oxidised in solution, the  $\text{SO}_3$  formed reacts with water to give  $\text{H}_2\text{SO}_4$ . The oxidising agent can be  $\text{Cl}_2$ ,  $\text{Br}_2$  or  $\text{H}_2\text{O}_2$ .

#### Oxidation by $\text{Cl}_2$ or $\text{Br}_2$ :

Sulphuric acid is formed when  $\text{SO}_2$  gas is passed through chlorine water or bromine water.



#### Oxidation of $\text{H}_2\text{O}_2$ :



If the acid is heavily contaminated, it is concentrated and thermally decomposed (to  $\text{H}_2\text{O}$ ,  $\text{SO}_2$  and  $\text{O}_2$ ) under oxidative or reductive conditions,

depending on the nature of the impurities. Oxidative process has the advantage that they liberate  $\text{CO}_2$ . Metal sulphates also provide a source of sulphuric acid. For example, gypsum or calcium sulphate (from the manufacture of phosphoric acid) can be reacted with coal; the calcium oxide is utilized for the manufacture of Portland cement



Alternatively, sulphuric acid can be prepared from the iron (II) sulphate waste, which is obtained during the manufacture of  $\text{TiO}_2$ .

The different methods of manufacturing sulphuric acid are essentially the same in principle and consist of three distinct processes :

- Production of sulphur dioxide
- Conversion of sulphur dioxide to sulphur trioxide
- Conversion of sulphur trioxide to sulphuric acid.

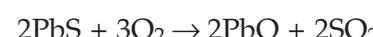
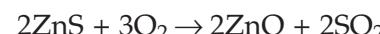
#### 1. Sulphur Dioxide is obtained in the following ways:

- (a) By burning elemental sulphur :



- (b) As a by-product of the roasting process in the extraction of certain metals from their sulphide ores.

#### For example :



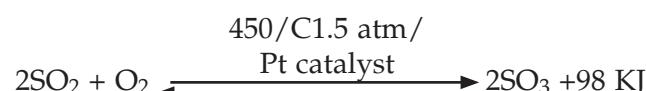
Since arsenic is often found in nature associated with sulphide ores, sulphur dioxide obtained by this method may contain some arsenic (III) oxide as impurity and in certain processes this is a distinct disadvantage.

### 8.19. LARGE SCALE PREPARATION

Commercially, sulphuric acid is produced by the contact process.

#### Principle

Sulphur dioxide is oxidised to sulphur trioxide by oxygen in the presence of a catalyst.



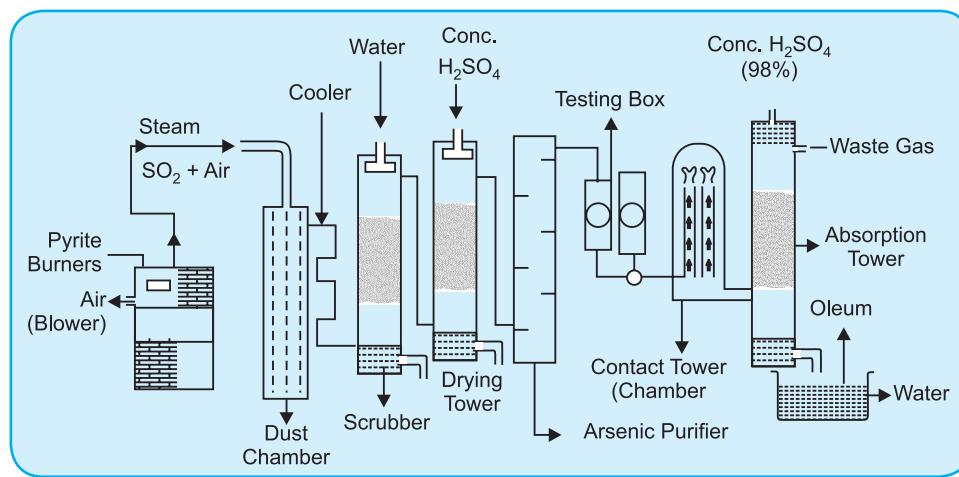


Fig. 8.18 : Preparation of Sulphuric Acid

### The Effect of Temperature

Applying Le Chatelier's principle, we can find out the effect of temperature on the equilibrium. The forward reaction, i.e., the formation of SO<sub>3</sub>, is exothermic, and the backward reaction, i.e., the decomposition of SO<sub>3</sub>, is endothermic. Therefore, if we raise the temperature of the equilibrium mixture, to counteract the effect of this change, SO<sub>3</sub> will absorb heat, and decompose into SO<sub>2</sub> and O<sub>2</sub>. Thus, the higher the temperature, the lower will be the yield of SO<sub>3</sub>. Conversely, the lower the temperature, the higher will be the yield of SO<sub>3</sub>. But, at lower temperatures, the reaction is slow. A temperature of 450°C is found to be appropriate for the reaction.

### The Effect of Pressure

From the equation, it is clear that three volumes of the reactants give two volumes of the product. Decrease in volume is favoured by an increase in pressure. Therefore, the higher the pressure, the greater is the yield of SO<sub>3</sub>. Generally, at a pressure of 1.5 atm, the yield is sufficient.

### Using A Catalyst

Platinum is used as a catalyst for the oxidation of SO<sub>2</sub> so that the equilibrium is attained fast. For various reasons, Vanadium pentoxide (V<sub>2</sub>O<sub>5</sub>) mixed with K<sub>2</sub>O (as promoter) on silica is preferred to Pt as a catalyst. As the oxidation takes place in 'contact' with the catalysts, the process is called the contact process.

### The Effect of the Removal of SO<sub>3</sub>

The removal of SO<sub>3</sub> from the equilibrium mixture will make the equilibrium shift in the forward direction. SO<sub>3</sub> does not dissolve smoothly in water to form H<sub>2</sub>SO<sub>4</sub>. Instead, it forms foggy droplets, which do not condense easily. However, it is smoothly

absorbed by 98% sulphuric acid to form pyrosulphuric acid (H<sub>2</sub>S<sub>2</sub>O<sub>7</sub>), a fuming liquid, also called oleum or fuming sulphuric acid.

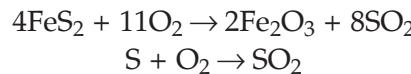


Oleum can be diluted by water to obtain sulphuric acid of the desired concentration.



### Method

SO<sub>2</sub> is obtained by roasting sulphide ores in pyrite burners.



It is mixed with air and passed through the catalyst chamber at 450°C and 1.5 atm. A yield of about 98-99% SO<sub>3</sub> is obtained and dissolved in concentrated H<sub>2</sub>SO<sub>4</sub> to obtain oleum. Oleum is diluted by water to get concentrated H<sub>2</sub>SO<sub>4</sub>. Commercial H<sub>2</sub>SO<sub>4</sub> is generally 98% acid.

## 8.20. PROPERTIES

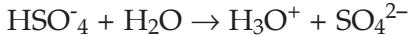
Table 8.2. Physical Properties of Sulphuric Acid

Physical Properties :	
1. Colour	Colourless
2. Odour	Odourless
3. Taste	Slightly sour in taste
4. Freezing point	10.371°C
5. Boiling point	290 - 317°C
6. Viscosity(25°C)	24.54Cp
7. Density(25°C)	1.8269 g cm <sup>-3</sup>
8. Solubility	Soluble in water in all proportion
9. Conductivity	In pure form it's non-conductor of electricity and in dilute form its good conductor of electricity.

## 1. Chemical Properties

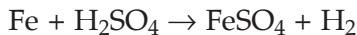
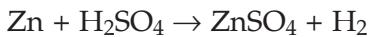
### 1. Acidic properties of dilute H<sub>2</sub>SO<sub>4</sub>.

H<sub>2</sub>SO<sub>4</sub> is a strong, dibasic acid. It undergoes primary dissociation and secondary dissociation.



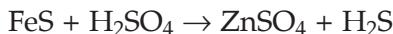
(a) Being a dibasic acid, the acid forms two series of salt – acid salts, e.g., sodium hydrogen sulphate (NaHSO<sub>4</sub>) and normal salts, e.g., sodium sulphate (Na<sub>2</sub>SO<sub>4</sub>).

(b) Dilute sulphuric acid liberates H<sub>2</sub> gas when it reacts with active metals.



(c) Reaction with carbonates, sulphites and sulphides

A dilute solution of the acid gives CO<sub>2</sub> with carbonates, SO<sub>2</sub> with sulphites and H<sub>2</sub>S with sulphides. The gases are liberated with effervescence.



## 2. Oxidising Properties

Sulphuric acid is a strong oxidising agent; while it is being reduced to SO<sub>2</sub>.

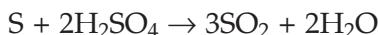
### Reaction with Metals

Metals react with hot concentrated H<sub>2</sub>SO<sub>4</sub> to give the metal sulphate and SO<sub>2</sub>.



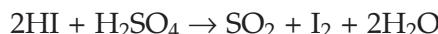
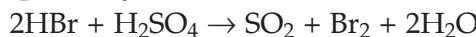
### Reactions with Non-Metals

Non-metals form their respective oxides when heated with concentrated H<sub>2</sub>SO<sub>4</sub>. For instance, C and S are oxidised to CO<sub>2</sub> and SO<sub>2</sub> respectively.



### Reaction with HBr and HI

Concentrated H<sub>2</sub>SO<sub>4</sub> oxidises HBr and HI to Br<sub>2</sub> and I<sub>2</sub> respectively.



## 3. Dehydrating Action of Concentrated H<sub>2</sub>SO<sub>4</sub>

Due to its great affinity towards water concentrated H<sub>2</sub>SO<sub>4</sub> functions not only as a drying agent, but also as a dehydrating agent. A dehydrating agent removes

the element of water, i.e., hydrogen and oxygen in stoichiometric ratio from a compound.

### (a) Dehydration of sugar

Sugar is a carbohydrate which has the general formula C<sub>x</sub>(H<sub>2</sub>O)<sub>y</sub>, for example, cane sugar (C<sub>12</sub>H<sub>22</sub>O<sub>11</sub>) may be represented as C<sub>12</sub>(H<sub>2</sub>O)<sub>11</sub>. Concentrated H<sub>2</sub>SO<sub>4</sub> dehydrates a sugar (or any carbohydrate) and converts it into black mass of carbon.

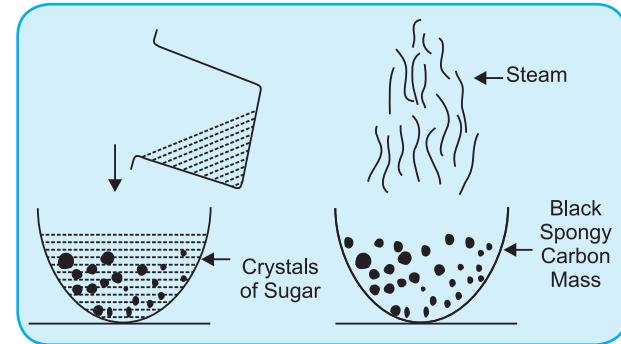
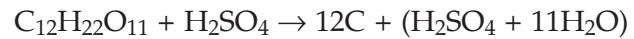
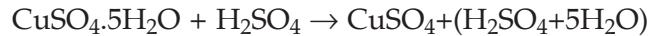


Fig. 8.19 : Dehydration of Sugar



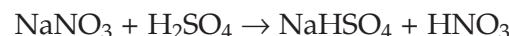
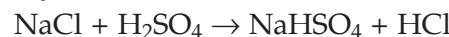
### (b) Dehydration of hydrated copper(II) Sulphate

Hydrated copper (II) Sulphate, CuSO<sub>4</sub>.5H<sub>2</sub>O loses its colour and turns into white anhydrous salt when kept in concentrated H<sub>2</sub>SO<sub>4</sub>.



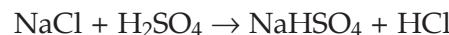
## 4. Displacement of a More Volatile Acid

It liberates HCl and HNO<sub>3</sub> from the salts due to its low volatility.



### Tests For H<sub>2</sub>SO<sub>4</sub>

1. It turns blue litmus into red.
2. Sulphuric acid reacts with any sulphate solution to give its corresponding sulphate product as a precipitate and this precipitate does not dissolve in conc. HCl or HNO<sub>3</sub>.
3. It gives a white precipitate of PbSO<sub>4</sub> with a solution of lead salts.
4. On heating with NaCl it gives out pungent fumes of HCl gas.



## 8.21. USES OF SULPHURIC ACID

Commercially, sulphuric acid has many important applications. The largest proportion is used in the manufacture of phosphorus and nitrogen based fertilizers. It is also used in the manufacture of inorganic sulphates, HF, B<sub>2</sub>O<sub>3</sub> and a range of organic

compounds, as well as in the petrochemical industry for the refining of raw paraffin.

1. Used as the laboratory reagent.
2. In the manufacture of other acids like HCl, HNO<sub>3</sub>, etc.
3. In the refining of petroleum.

4. In the storage of batteries.
5. For cleansing metal surfaces before galvanization.
6. In the textile, paper and dyeing industries.
7. In the manufacture of explosives e.g., nitroglycerine, trinitrotoluene(TNT), etc.
8. In the manufacture of medicines and disinfectants.
9. As a dehydrating and desiccating agent.

## INTEXT QUESTIONS

1. Give the electron dot diagram and orbital diagram of HCl.
2. Give chemical properties with reactions for hydrogen chloride gas.
3. Give the reaction for the formation of Aqua Regia.
4. Give the laboratory preparation of ammonia.
5. Give the brown ring test with reactions.
6. Give the oxidising property of H<sub>2</sub>SO<sub>4</sub>.
7. Give the chemical properties of sulphuric acid with complete reactions.
8. Give the fountain experiment to demonstrate the high solubility of HCl in water.
9. Give the action HCl on the salts of weaker acids with complete reactions.
10. Give the method of preparation of aqueous ammonia in the laboratory.



## POINTS TO REMEMBER

- ◆ HCl gas is prepared by the photochemical combination of H<sub>2</sub> and Cl<sub>2</sub>, action of concentrated H<sub>2</sub>SO<sub>4</sub> on metal chlorides NH<sub>4</sub>Cl, and chlorination of hydrocarbons.
- ◆ It is a colourless gas and is highly soluble in water as demonstrated by the fountain experiment.
- ◆ HCl gas is oxidised by strong oxidising agents to Cl<sub>2</sub>.
- ◆ An aqueous solution of HCl is called hydrochloric acid.
- ◆ On a large scale, hydrochloric acid is prepared by first strongly heating NaCl with concentrated H<sub>2</sub>SO<sub>4</sub> (the Leblanc process or the salt-cake process). The HCl gas evolved is dissolved in water in an absorption tower.
- ◆ The commercial hydrochloric acid is often yellow due to the dissolved FeCl<sub>3</sub> as impurity.
- ◆ The constant-boiling acid has 20.24%, HCl by mass and boils at 110°C.
- ◆ The electrolysis of dilute HCl acid gives H<sub>2</sub> at the cathode and O<sub>2</sub> at the anode and that of concentrated HCl gives H<sub>2</sub> at the cathode and Cl<sub>2</sub> at the anode.
- ◆ Hydrochloric acid reacts with metals above hydrogen in the activity series to liberate H<sub>2</sub>.
- ◆ It neutralises metal oxides and hydroxides to form the respective chlorides.
- ◆ Hydrochloric acid gives a white precipitate of AgCl with an AgNO<sub>3</sub> solution.
- ◆ NH<sub>3</sub> is prepared by
  - (i) Heating ammonium salts with alkalis
  - (ii) The hydrolysis of metal nitrides
  - (iii) The direct combination of N<sub>2</sub> and H<sub>2</sub> (Haber process)
- ◆ NH<sub>3</sub> is prepared in the laboratory by heating ammonium salts with slaked lime.
- ◆ NH<sub>3</sub> is dried over quicklime (CaO). Common drying agents like concentrated H<sub>2</sub>SO<sub>4</sub>, anhydrous CaCl<sub>2</sub> or P<sub>2</sub>O<sub>5</sub> cannot be used because they react with NH<sub>3</sub>. Being lighter than air, the gas is collected by the downward displacement of air. NH<sub>3</sub> is manufactured on a large scale by the direct combination of N<sub>2</sub> and H<sub>2</sub> in Haber process.
- ◆ Ammonia is separated from the reactants in Haber process by liquefaction. Ammonia is a colourless gas with a characteristic smell. It turns moist red litmus blue, showing its basic character.
- ◆ Ammonia is highly soluble in water, as demonstrated by the fountain experiment.



- ◆ Ammonia reacts with acids to form salts. Gaseous ammonia reacts with gaseous HCl to form white fumes of  $\text{NH}_4\text{Cl}$ .
- ◆ Ammonia is a reducing agent.
- ◆ Nitric acid is prepared by distilling  $\text{KNO}_3$  or  $\text{NaNO}_3$  with concentrated  $\text{H}_2\text{SO}_4$ .
- ◆  $\text{HNO}_3$  decomposes upon heating or on exposure to light to form brown  $\text{NO}_2$  gas.
- ◆ The pure acid is colourless but often it is yellow or even red owing to the presence of some dissolved oxides of nitrogen. These oxides may be removed by bubbling dry air through the acid at 60–80°C.
- ◆ Fuming  $\text{HNO}_3$  contains dissolved oxides of nitrogen. It is prepared by distilling  $\text{NaNO}_3$  or  $\text{KNO}_3$  with concentrated  $\text{H}_2\text{SO}_4$  in the presence of starch.
- ◆  $\text{HNO}_3$  is manufactured on a large scale by the Ostwald process, which is based on the catalytic oxidation of  $\text{NH}_3$ .
- ◆  $\text{HNO}_3$  forms nitrate salts with metal oxides, hydroxides, carbonates and hydrogen carbonates.
- ◆ Aqua regia is a mixture of concentrated  $\text{HNO}_3$  1 part, and concentrated HCl 3 parts.
- ◆ Noble metals like gold and platinum dissolve only in the aqua regia mixture.
- ◆  $\text{HNO}_3$  is used in the manufacture of fertilizers and explosives.
- ◆ Sulphuric acid is formed by the dissolution of  $\text{SO}_3$  in water, and the oxidation of moist  $\text{SO}_2$

- by chlorine water, bromine water or hydrogen peroxide.
- ◆  $\text{H}_2\text{SO}_4$  is manufactured on a large scale by the contact process, which involves the conversion of  $\text{SO}_2$  to  $\text{SO}_3$  by oxygen in the presence of platinum as a catalyst at the optimum temperature of 450°C and a pressure of 1.5 atm.
- ◆ The absorption of  $\text{SO}_3$  into a 98%  $\text{H}_2\text{SO}_4$  solution to form fuming sulphuric acid, or oleum ( $\text{H}_2\text{S}_2\text{O}_7$ ).
- ◆ The dilution of oleum by water to obtain  $\text{H}_2\text{SO}_4$  of the desired concentration.
- ◆  $\text{SO}_3$  is not completely absorbed in water due to the formation of a dense fog of minute droplets of  $\text{H}_2\text{SO}_4$ .
- ◆ The commercial acid has a concentration of 98%.
- ◆ Pure  $\text{H}_2\text{SO}_4$  is a colourless, oily liquid.
- ◆ Concentrated  $\text{H}_2\text{SO}_4$  is much less volatile than either  $\text{HNO}_3$  or HCl. Therefore, it displaces HCl from chlorides and  $\text{HNO}_3$  from nitrates.
- ◆  $\text{H}_2\text{SO}_4$  is a dibasic acid forming two series of salts: the acid hydrogen sulphates and the normal sulphates.
- ◆ The dilute acid gives  $\text{H}_2$  with metals above hydrogen in the activity series.
- ◆ Dilute  $\text{H}_2\text{SO}_4$  gives  $\text{CO}_2$ ,  $\text{SO}_2$  and  $\text{H}_2\text{S}$  respectively with carbonates, sulphites and sulphides.
- ◆ Hot and concentrated  $\text{H}_2\text{SO}_4$  acts as an oxidising agent. It oxidises C to  $\text{CO}_2$ , and S to  $\text{SO}_2$ .
- ◆ Sulphuric acid is of great commercial importance.

## EXERCISES

### A. FILL IN THE BLANKS

1. HCl was prepared in 1648 by .....
2. HCl gas has the .....odour.
3.  $\text{NH}_3$  molecule has ..... one pair of electron.
4. Aqua Regia is a mixture of 3:1, HCl and .....
5.  $\text{H}_2\text{SO}_4$  turns blue litmus .....
6.  $\text{H}_2\text{SO}_4$  gives ..... when reacts with  $\text{PbSO}_4$ .
7. Quicklime is not used to dry HCl gas because ..... . [2014]
8. Ammonia gas is collected by ..... . [2014]
9. Cold, dilute nitric acid reacts with copper to form ..... . [2014]

### B. MULTIPLE CHOICE TYPE QUESTIONS

Choose the correct answer from the options given below :

1. Fountain experiment was used to demonstrate :  
(a) High solubility    (b) Density  
(c) Low solubility    (d) Vapour pressure
2. Methyl orange turns HCl from orange to  
(a) Pink                (b) Yellow  
(c) Green              (d) Blue
3. Aqua Regia is a mixture of 3:1  
(a)  $\text{HCl} : \text{HNO}_3$     (b)  $\text{HNO}_3 : \text{HCl}$   
(c)  $\text{H}_2\text{SO}_4 : \text{H}_2\text{SO}_4$     (d)  $\text{HNO}_3 : \text{H}_2\text{SO}_4$
4. Complete the following Reaction :  
 $2\text{NH}_3 + 3 \text{Cl}_2 \rightarrow \underline{\hspace{2cm}} + 6\text{HCl}$

- |                    |                        |
|--------------------|------------------------|
| (a) N <sub>2</sub> | (b) NH <sub>2</sub> Cl |
| (c) H <sub>2</sub> | (d) 2N <sub>2</sub>    |
5. Brown ring test is carried out for :
- |              |                   |
|--------------|-------------------|
| (a) Nitrate  | (b) Oxalate       |
| (c) Sulphate | (d) Cyanide group |

### C. ANSWER THE FOLLOWING QUESTIONS BASED ON THE GIVEN INSTRUCTIONS

- (1) Identify the acid which matches the following description (i) to (v) :
- The acid which is used in the preparation of a non-volatile acid.
  - The acid which produces sugar charcoal from sugar.
  - The acid which is prepared by catalytic oxidation of ammonia.
  - The acid on mixing with lead nitrate solution produces a white precipitate which is insoluble even on heating.
  - The acid on mixing with silver nitrate solution produces a white precipitate which is soluble in excess ammonium hydroxide. [ICSE, 2015]
- (2) Select from the list the gas that matches the description given in each case :
- [ammonia, ethane, hydrogen chloride, hydrogen sulphide, ethyne]
- This gas is used as a reducing agent in reducing copper oxide to copper.
  - This gas produces dense white fumes with ammonia gas. [ICSE, 2015]

### D. MATCH THE COLUMNS

Column I	Column II
1. Fountain experiment	HNO <sub>3</sub>
2. Xanthophyll	Ferrous sulphate
3. Sour in taste	H <sub>2</sub> SO <sub>4</sub>
4. Brown ring test	Protein (yellow colour)
5. Manufacture of dyes	High solubility

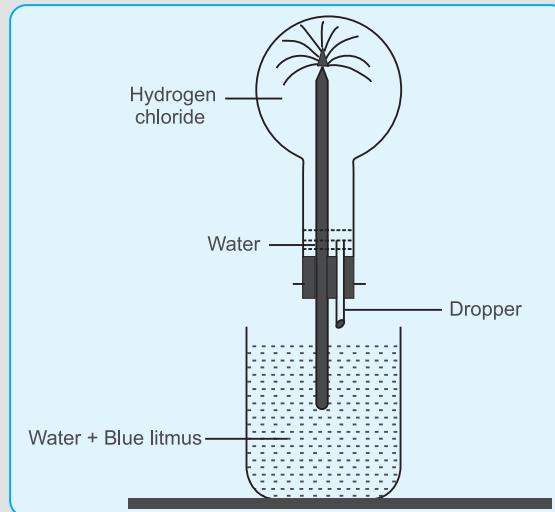
### E. GIVE THE BALANCED CHEMICAL EQUATIONS FOR THE FOLLOWING

- Thermal dissociation of HCl gas.
- Reaction of ammonia with HCl gas.
- Reaction of HCl with sodium atom.
- Reaction of a mixture of hydrochloric acid and nitric acid.
- Reaction of ammonia with water molecule.
- To show oxidising property of H<sub>2</sub>SO<sub>4</sub>.

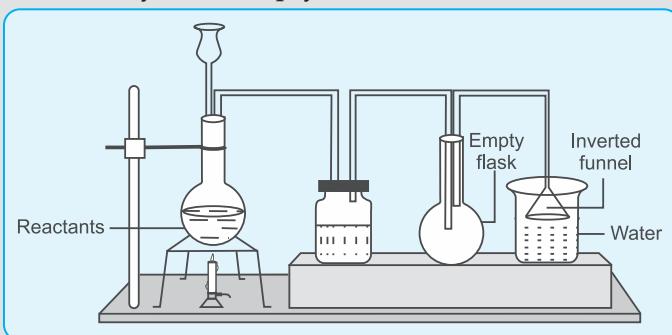
### F. VERY SHORT QUESTIONS

- Write balanced equation.
  - Copper oxide and dilute hydrochloric acid.

- Managanese (iv) oxide and concentrated hydrochloric acid
- Name the experiment illustrated below.

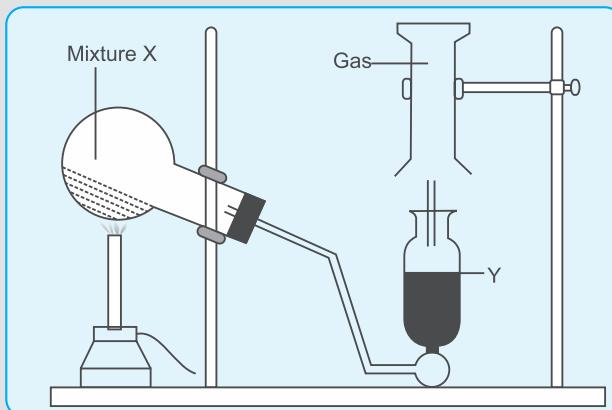


- Which property of hydrogen chloride is demonstrated by this experiment?
  - State the colour of the water that has entered the round-bottomed flask.
  - Why cannot HCl be dried over quicklime or P<sub>4</sub>O<sub>10</sub>?
  - Why is commercial hydrochloric acid often yellow in colour ?
  - The given set up in the figure is for the preparation of an acid.
- Name the acid prepared by this method.
  - Name the reactants used.
  - Why is an empty flask used ?



- Name the experiment, which demonstrates that HCl is highly soluble in water. [ICSE 2000]
- Why is an aqueous solution of HCl a good conductor of electricity ?
- Which among Cu, Fe, Mg, Na and Zn does not react with a dilute HCl solution ? [ICSE 2001]
- A solution of hydrgoen chloride in water is prepared. Separate portions of the solution are treated with the following substance

- (a) Calcium carbonate  
 (b) Magnesium ribbon  
 (c) Manganese (IV) oxide with heating  
 (d) Sodium sulphide
- Name the gas evolved in each case and its odour.  
 [ICSE 2004]
9. Name the gas obtained by warming ammonium sulphate with sodium hydroxide solution.
10. Write the equation for the formation of ammonia by the action of water on magnesium nitride.  
 [ICSE 2003]
11. What are the products formed when ammonia is oxidised by copper oxide ?  
 [ICSE 2001]
12. What are the compounds required for the laboratory preparation of nitric acid ?  
 [ICSE 2002]
13. How is dilute  $\text{HNO}_3$  solution concentrated ?
14. Containers of which metal can be used to store concentrated  $\text{HNO}_3$ .  
 [ICSE 2002]
15. Which gas is produced when potassium nitrate is heated ? Write the equation for the reaction.  
 [ICSE 2003]
16. Name the gas produced (i) by the action of dilute nitric acid on copper, and (ii) on heating sodium nitrate.  
 [ICSE 2009]
17. What is the property of nitric acid, which allows it to react with copper ?  
 [ICSE 2008]
18. The diagram shows an experimental set up for the laboratory preparation of a pungent smelling gas. The gas is alkaline in nature.  
 [ICSE 2011]

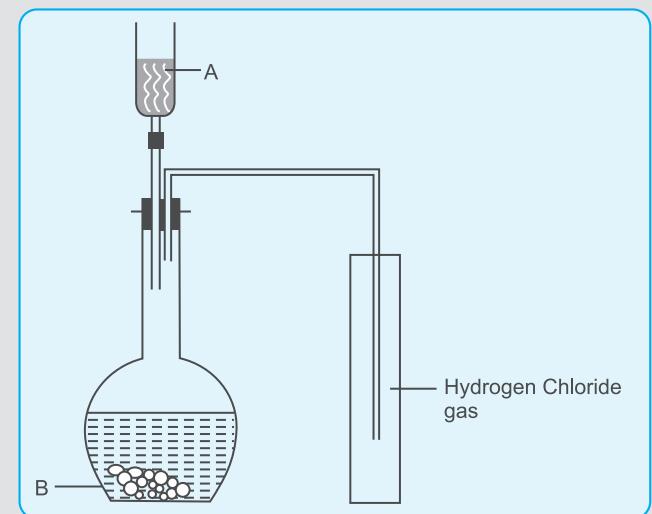


- (i) Name the gas collected in the jar.  
 (ii) Write the balanced equation for the above preparation.  
 (iii) How is the gas being collected ?  
 (iv) Name the drying agent used.  
 (v) How will you find that the jar is full of gas ?

19. Name two oxidising agents that will oxidise  $\text{SO}_2$  into  $\text{H}_2\text{SO}_4$ .
20. Name the catalyst used industrially, which speeds up the conversion of  $\text{SO}_2$  into  $\text{SO}_3$ . [ICSE 2000]
21. What is the name of the compound formed between sulphur trioxide and sulphuric acid ?  
 [ICSE 2000]
22. What is the property of concentrated sulphuric acid, which allows it to be used in the preparation of hydrogen chloride and nitric acid ? [ICSE 2008]
23. Give two examples of  $\text{H}_2\text{SO}_4$  acting as an oxidising agent.
24. Name the process used for the large scale manufacture of sulphuric acid.
25. Which property of sulphuric acid; accounts for its use as a dehydrating agent?

### G. SHORT QUESTIONS

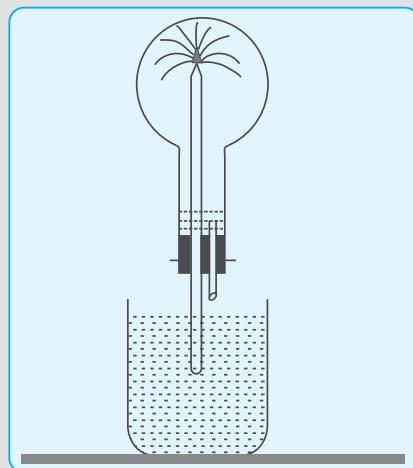
1. What must be added to  $\text{NaCl}$  to obtain  $\text{HCl}$  ? Write the equation for the reaction which takes place.  
 [ICSE 2000]
2. What is Azeotropic mixture ?
3. The diagram shows an apparatus for the laboratory preparation of hydrogen chloride.  
 [ICSE 2010]



Identify A and B

- (i) Write the equation for the reaction.  
 (ii) How would you check whether the gas jar is filled with hydrogen chloride ?  
 (iii) What does the method of collection tell you about the density of hydrogen chloride ?
4. Write balanced equations for the reaction between
- (a) Dilute hydrochloric acid and Sodium sulphite  
 [ICSE 2001]
- (b) Hydrogen chloride and Ammonia  
 [ICSE 2004, 2009]

- (c) Red lead and concentrated Hydrochloric acid [ICSE 2005]
- (d) Copper oxide and dilute Hydrochloric acid [ICSE 2009]
- (e) Dilute hydrochloric acid and Magnesium metal [ICSE 2009]
5. In the laboratory preparation of hydrochloric acid, HCl gas is dissolved in water.
- Draw a diagram to show the arrangement used for the absorption of HCl in water.
  - Why is such an arrangement necessary? Give two reasons.
  - Write the chemical equations for the laboratory preparation of HCl gas when the reactants are : (A) below 200°C (B) above 200°C [ICSE 2011]
6. What happens when dilute hydrochloric acid is added to a lead nitrate solution ? [ICSE 2002]
7. What is pickling ?
8. Write balanced equations for the following reactions.
- A mixture of ammonium chloride and slaked lime is heated [ICSE 2005]
  - Aluminium nitride is heated with water [ICSE 2005]
9. (a) Write the equation for the reaction in the Haber process for the large-scale manufacture of ammonia. [ICSE 2004]
- (b) Name the catalyst used in the Haber synthesis of NH<sub>3</sub>.
- (c) State the purpose of liquefying the ammonia produced in the process. [ICSE 2004]
10. Which compound is normally used as a drying agent for ammonia ? [ICSE 2003]
11. Why are hydrous CaCl<sub>2</sub>, P<sub>2</sub>O<sub>5</sub> or concentrated H<sub>2</sub>SO<sub>4</sub> not used for drying NH<sub>3</sub> ?
12. How is ammonia collected ? [ICSE 2003]
13. Why is ammonia not collected over water ? [ICSE 2003]
14. Which feature of the ammonia molecule leads to the formation of the ammonium ion when ammonia dissolves in water ?
15. Write the balanced equations for the reaction of
- Hydrogen chloride with ammonia [ICSE 2001, 2004]
  - Chlorine with an excess of ammonia [ICSE 2004]
16. (a) When nitric acid is prepared by the action of concentrated sulphuric acid on potassium nitrate, what is the special feature of the apparatus used ? [ICSE 2003]
- (b) Write the equation for the laboratory preparation of nitric acid from potassium nitrate and concentrated sulphuric acid. [ICSE 2003]
17. During a thunderstorm, the rainwater contains HNO<sub>3</sub>. The HNO<sub>3</sub> is formed as a result of three chemical reaction. Write the balanced chemical equations. [ICSE 1995]
18. Account for the yellow colour that appears in concentrated nitric acid when it is left standing in an ordinary glass bottle. [ICSE 2005]
19. Write balanced equations for the following reactions.
- A mixture of sodium nitrate and concentrated sulphuric acid is heated. [ICSE 2005]
  - Concentrated nitric acid is added to copper turnings kept in a beaker. [ICSE 2004]
  - Dilute nitric acid is added to copper. [ICSE 2005]
  - Sulphur is treated with hot concentrated nitric acid. [ICSE 2005]
  - Dilute nitric acid is added to copper (II) oxide. [ICSE 2002]
  - Solid lead nitrate is heated in a dry test tube.
20. Explain why only an all-glass apparatus should be used for the preparation of nitric acid by heating concentrated sulphuric and potassium nitrate.
21. The questions below are related to the manufacture of ammonia. [ICSE 2010]
- Name the process.
  - In what ratio must the reactants be taken ?
  - Name the catalyst used.
  - Give the equation for the manufacture of ammonia.
  - Ammonia can act as a reducing agent-write a relevant equation for such a reaction.
22. The diagram shows a simple arrangement of the fountain experiment : [ICSE 2010]



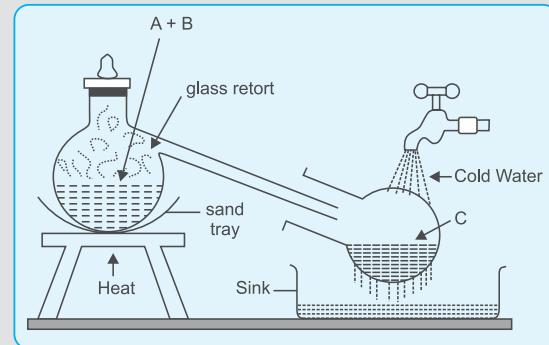
- (a) Name the two gases you have studied which can be used in this experiment.
- (b) What is the common property demonstrated by this experiment?
- (c) The action of heat on the blue crystalline solid L gives a reddish brown gas M, a gas that relights a glowing splint and leaves a black residue. When gas N, which has a rotten egg smell, is passed through a solution of L a black precipitate is formed.
- Identify L, M and N (Name or formula).
  - Write the equation for the action of heat on L.
- (d) Write the equation for the reaction between the solution of L and the gas N.
23. Write a chemical equation to illustrate the acidic nature of nitric acid. [ICSE 2006]
24. How is sulphur converted into sulphuric acid?
25. Why is a direct reaction between  $\text{SO}_3$  and water avoided in the contact process?
26. Give the balanced equation of the roasting of iron pyrites.
27. Write balanced equations for the reactions in which the following gases are obtained using dilute  $\text{H}_2\text{SO}_4$  as one of the reactant. (i)  $\text{H}_2$  (ii)  $\text{CO}_2$  (iii)  $\text{SO}_2$  [ICSE 1995]
28. By what means other than heating can you dehydrate  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  and obtain anhydrous copper sulphate. [ICSE 1998]
29. What is oleum? Why it is also called fuming sulphuric acid?
30. Concentrated sulphuric acid is both an oxidising agent and a non-volatile acid. Write one equation each to illustrate the above-mentioned properties of sulphuric acid. [ICSE 2006]

## H. LONG QUESTIONS

1. What is the special arrangement required for dissolving  $\text{HCl}$  gas in water in the laboratory

preparation of hydrochloric acid? Explain with a diagram the reason for such an arrangement.

- How is ammonia prepared in the laboratory?
- What are the experimental conditions that favour the formation of  $\text{NH}_3$  in the Haber synthesis from  $\text{N}_2$  and  $\text{H}_2$ ?
- The figure given below illustrates the apparatus used in the laboratory preparation of nitric acid.
- Write a short note on the oxidising properties of nitric acid.



- Name A (a liquid), B (a solid) and C (a liquid). (Do not give the formulae).
- Write an equation to show how nitric acid undergoes decomposition.
- Write the equation for the reaction in which copper is oxidised by concentrated nitric acid. [ICSE 2007]
- Discuss the Ostwald process for the manufacture of nitric acid from ammonia.
- Identify the following substances: [ICSE 2012]
  - An acidic gas which gives dense white fumes with  $\text{NH}_3$ .
  - A solid which when kept in the open forms a solution after sometime.
  - A metal which gives hydrogen gas on reacting with both dilute acid and alkali.
- How is  $\text{HNO}_3$  prepared in the laboratory?

## RECENT BOARD QUESTIONS

### Question 1.

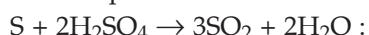
- (a) Choose the correct answer from the options given below : [2011, 13, 16]

- (i) Hydrogen chloride gas being highly soluble in water is dried by :
- Anhydrous calcium chloride
  - Phosphorous penta oxide
  - Quick lime
  - Concentrated sulphuric acid

- The brown ring test is used for detection of :
  - $\text{CO}_3^{2-}$
  - $\text{NO}_3^-$
  - $\text{SO}_3^{2-}$
  - $\text{Cl}^-$
- When dilute sulphuric acid reacts with iron sulphide, the gas evolved is .....
  - Hydrogen sulphide
  - Sulphur dioxide
  - Sulphur trioxide
  - Vapour of sulphuric acid



- (iv) In the given equation identify the role played by concentrated sulphuric acid



- (A) Non-volatile acid
- (B) Oxidising agent
- (C) Dehydrating agent
- (D) None of the above.

- (v) Nitrogen gas can be obtained by heating :

- (A) Ammonium nitrate
- (B) Ammonium nitrite
- (C) Magnesium nitride
- (D) Ammonium chloride.

- (vi) The aim of the **fountain experiment** is to prove that :

- (A) HCl turns blue litmus red
- (B) HCl is denser than air
- (C) HCl is highly soluble in water
- (D) HCl fumes in moist air

**(b) What would you observe in each of the following cases :** [2011,13,15,16]

- (i) Sugar crystals are added to a hard glass test tube containing concentrated sulphuric acid.
- (ii) Copper is heated with concentrated nitric acid in a hard glass test tube.
- (iii) When carbon monoxide is passed over heated copper oxide.
- (iv) Concentrated sulphuric acid is added drop wise to a crystal of hydrated copper sulphate.
- (v) Copper sulphide is treated with dilute hydrochloric acid.
- (vi) Excess of chlorine gas is reacted with ammonia gas.
- (vii) A few drops of dilute hydrochloric acid are added to silver nitrate solution, followed by addition of ammonium hydroxide solution.
- (viii) When hydrogen sulphide gas is passed through lead acetate solution.
- (ix) When ammonia gas is burnt in an atmosphere of excess oxygen.
- (x) Dilute hydrochloric acid is added to lead nitrate solution and the mixture is heated.

**(c) Answer the following questions :** [2011,15,16]

- (i) An aqueous salt solution used for testing sulphate radical.
- (ii) A catalyst used in the manufacture of nitric acid by Ostwald's process.
- (iii) This gas is used as a reducing agent in reducing copper oxide to copper.
- (iv) This gas produces dense white fumes with ammonia gas.

- (v) This gas has a characteristic rotten egg smell.
- (vi) Concentrated sulphuric acid is added to sugar crystals.
- (vii) Dilute hydrochloric acid is added to copper carbonate.
- (viii) Dilute hydrochloric acid is added to sodium thiosulphate.

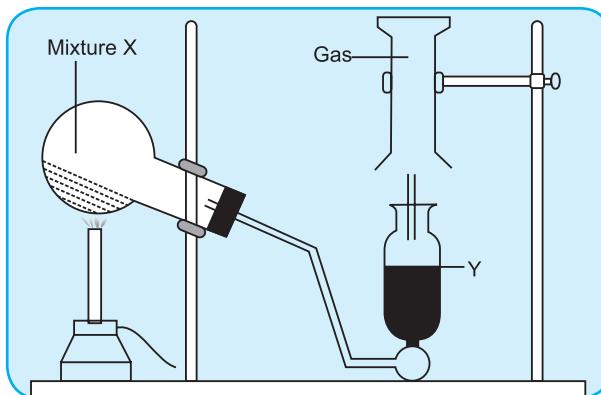
**(d) Write the balanced chemical equation for each of the following reactions :** [2011,12,13,14,15,16]

- (i) Sodium thiosulphate is reacted with dilute hydrochloric acid.
- (ii) Calcium bicarbonate reacts with dilute hydrochloric acid.
- (iii) Dilute sulphuric acid is poured over sodium sulphite.
- (iv) Chlorine reacts with excess of ammonia.
- (v) Ferric hydroxide reacts with nitric acid.
- (vi) Dilute nitric acid and Copper carbonate.
- (vii) Concentrated hydrochloric acid and Potassium permanganate solution.
- (viii) Ammonia and oxygen in the presence of a catalyst.
- (ix) Reduction of hot Copper (II) oxide to copper using ammonia gas.
- (x) Oxidation of carbon with concentrated nitric acid.
- (xi) Dehydration of concentrated sulphuric acid with sugar crystals.
- (xii) Action of heat on a mixture of copper and concentrated nitric acid.
- (xiii) Action of concentrated sulphuric acid on carbon.
- (xiv) Action of dilute hydrochloric acid on sodium sulphide.
- (xv) Lead sulphate from lead carbonate.
- (xvi) Sodium sulphate using dilute sulphuric acid.
- (xvii) Copper chloride using copper carbonate.
- (xviii) Lab preparation of ammonia using an ammonium salt.
- (xix) Reaction of ammonia with excess chlorine.
- (xx) Reaction of ammonia with sulphuric acid.
- (xxi) Action of hot and concentrated nitric acid on copper.
- (xxii) Action of hydrochloric acid on sodium bicarbonate.
- (xxiii) Action of dilute Sulphuric acid on Sodium Sulphite.
- (xxiv) When excess of ammonia is treated with chlorine.
- (xxv) An equation to illustrate the reducing nature of ammonia.

**Question 2.**

- (a) The diagram shows an experimental set up for the**

laboratory preparation of a pungent smelling gas. The gas is alkaline in nature. [2011]



- Name the gas collected in the jar.
  - Write the balanced equation for the above preparation.
  - How is the gas being collected?
  - Name the drying agent used.
  - How will you find that the jar is full of gas?
- (b) In the laboratory preparation of hydrochloric acid, HCl gas is dissolved in water. [2011]
- Draw a diagram to show the arrangement used for the absorption of HCl in water.
  - Why is such an arrangement necessary? Give two reasons.
  - Write the chemical equations for the laboratory preparation of HCl gas when the reactants are:
    - (A) below 200°C
    - (B) above 200°C

#### Question 3.

- (i) With the help of equations, give an outline for the manufacture of sulphuric acid by the contact process.  
 (ii) What property of sulphuric acid is shown by the reaction of concentrated sulphuric acid when heated with
  - (A) potassium nitrate (B) carbon
- (i) What is the special feature of the apparatus that is used in the laboratory preparation of nitric acid?  
 (ii) Why should the temperature of the reaction mixture of nitric acid not be allowed to rise above 200°C?

#### Question 4.

- Concentrated nitric acid oxidises phosphorus to phosphoric acid according to the following equation : [2012]  

$$P + 5HNO_3(\text{conc.}) \rightarrow H_3PO_4 + H_2O + 5NO_2$$
 If 9.3g of phosphorus was used in the reaction, calculate :
  - Number of moles of phosphorus taken.
  - The mass of phosphoric acid formed.

- The volume of nitrogen dioxide produced at STP.  
 $[H = 1, N = 14, P = 31, O = 16]$

- (b) Give reasons for the following : [2012]
- Iron is rendered passive with fuming nitric acid.
  - Hydrogen chloride gas cannot be dried over quick lime.

#### Question 5.

- (a) Some properties of sulphuric acid are listed below. Choose the role played by sulphuric acid as A, B, C or D which is responsible for the reactions (i) to (v) [2012]
- |                       |                       |
|-----------------------|-----------------------|
| (A) Dilute acid       | (B) Dehydrating agent |
| (C) Non-volatile acid | (D) Oxidising agent.  |
- $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} \xrightarrow{\text{conc. H}_2\text{SO}_4} \text{CuSO}_4 + 5\text{H}_2\text{O}$
  - $\text{S} + \text{H}_2\text{SO}_4(\text{conc.}) \rightarrow 3\text{SO}_2 + 2\text{H}_2\text{O}$
  - $\text{NaNO}_3 + \text{H}_2\text{SO}_4(\text{conc.}) \xrightarrow{200^\circ\text{C}} \text{NaHSO}_4 + \text{HCl}$
  - $\text{MgO} + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2\text{O}$
  - $\text{Zn} + 2\text{H}_2\text{SO}_4(\text{conc.}) \rightarrow \text{ZnSO}_4 + \text{SO}_2 + 2\text{H}_2\text{O}$

- (b) State one chemical test between each of the following pairs : [2012]

- Sodium carbonate and Sodium sulphite.
- Manganese dioxide and Copper (II) oxide.

- (c) Draw an electron dot diagram to show the structure of hydronium ion. State the type of bonding present in it. [2012]

- (d) The following questions are based on the preparation of ammonia gas in the laboratory : [2012]

- Explain why ammonium nitrate is not used in the preparation of ammonia.
- Name the compound normally used as a drying agent during the process.
- How is ammonia gas collected?
- Explain why it is not collected over water

#### Question 6.

- (a) Choosing the substances from the list given below, write balanced chemical equations for the reactions which would be used in the laboratory to obtain the following salts : [2013]

Dilute Sulphuric acid	Copper Iron	Copper(II) carbonate Sodium carbonate
Sodium	Sodium chloride	Zinc nitrate

- Sodium sulphate
- Zinc carbonate
- Copper(II) sulphate
- Iron(II) sulphate.

- (b) Identify the gas evolved in the following reactions when : [2013]

- Sulphur is treated with concentrated nitric acid.
- Concentrated hydrochloric acid is made to react with manganese dioxide

**Question 7.**

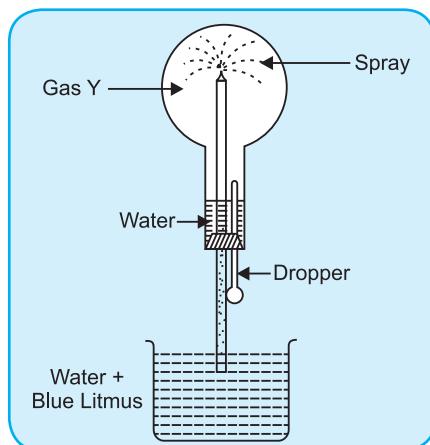
(a) Identify the following substances which are underlined : [2013]

- An alkaline gas which produces dense white fumes when reacted with hydrogen chloride gas.
- An acid which is present in vinegar.
- A gas which does not conduct electricity in the liquid state but conducts electricity when dissolved in water.
- A dilute mineral acid which forms a white precipitate when treated with barium chloride solution.
- The element which has the highest ionization potential.

(b) Distinguish between the following pairs of compounds using the test given within brackets : [2014]

- Iron (II) sulphate and iron (III) sulphate (using ammonium hydroxide).
- A lead salt and a zinc salt (using excess ammonium hydroxide).
- Sodium nitrate and sodium sulphite (using dilute sulphuric acid).
- Dilute sulphuric acid and dilute hydrochloric acid (using barium chloride solution).
- Ethane and ethene (using alkaline potassium permanganate solution).

(d) Study the figure given below and answer the questions that follow : [2014]



- Identify the gas Y.
  - What property of gas Y does this experiment demonstrate ?
  - Name another gas which has the same property and can be demonstrated through this experiment.
- (c) (i) Name the other ion formed when ammonia dissolves in water.
- (ii) Give one test that can be used to detect the presence of the ion produced.

**Question 8.**

(a) State the conditions required for the following reactions to take place : [2014]

- Catalytic oxidation of ammonia to nitric oxide.
- Any two conditions for the conversion of sulphur dioxide to sulphur trioxide.

(b) Give balanced equations for the Laboratory preparation of nitric acid. [2014]

(c) Give one equation each to show the following properties of sulphuric acid: [2014]

- Dehydrating property.
- Acidic nature.
- As a non-volatile acid.

(d) Identify the acid which matches the following description (i) to (v) : [2015]

- The acid which is used in the preparation of a non-volatile acid.
- The acid which produces sugar charcoal from sugar.
- The acid which is prepared by catalytic oxidation of ammonia.
- The acid on mixing with lead nitrate solution produces a white precipitate which is insoluble even on heating.

**Question 9.**

(a) Identify the anion present in each of the following compounds : [2015]

- A salt M on treatment with concentrated sulphuric acid produces a gas which fumes in moist air and gives dense fumes with ammonia.
- A salt D on treatment with dilute sulphuric acid produces a gas which turns lime water milky but has no effect on acidified potassium dichromate solution.

(b) (i) Give balanced chemical equations for the action of sulphuric acid on each of the following :

- Potassium hydrogen carbonate.
- Sulphur

(ii) In the contact process for the manufacture of sulphuric acid give the equations for the conversion of sulphur trioxide to sulphuric acid.

(c) Explain the following : [2015]

- Dilute nitric acid is generally considered a typical acid but not so in its reaction with metals.
- Concentrated nitric acid appears yellow when it is left standing in a glass bottle.
- An all glass apparatus is used in the laboratory preparation of nitric acid.



(d) The following questions are pertaining to the laboratory preparation of hydrogen chloride gas :

[2015]

- (i) Write the equation for its preparation mentioning the condition required.
- (ii) Name the drying agent used and justify your choice.
- (iii) State a safety precaution you would take during the preparation of hydrochloric acid.

**Question 10.**

(a) Identify the gas evolved and give the chemical test in each of the following cases : [2016]

- (i) Dilute hydrochloric acid reacts with sodium sulphite.

(ii) Dilute hydrochloric acid reacts with iron (II) sulphide.

(b) A, B, C and D summarize the properties of sulphuric acid depending on whether it is dilute or concentrated.

[2016]

A = Typical acid property

B = Non-volatile acid

C = Oxidizing agent

D = Dehydrating agent

Choose the property (A, B, C or D) depending on which is relevant to each of the following :

- (i) Preparation of Hydrogen chloride gas.
- (ii) Preparation of Copper sulphate from copper oxide.
- (iii) Action of conc. Sulphuric acid on Sulphur.



# 9

# ORGANIC CHEMISTRY

## LEARNING OUTCOMES

- 9.0. Introduction
- 9.1. Comparison between Organic and Inorganic Compounds
- 9.2. Application of Organic Chemistry
- 9.3. Unique Nature of Carbon Atoms
- 9.4. Types of Organic Compounds
- 9.5. Classification of Organic Compounds
- 9.6. Different Types of Structural Isomerism
- 9.7. Homologous Series
- 9.8. Characteristics of Homologous Series
- 9.9. Nomenclature
- 9.10. Hydrocarbons
- 9.11. Alkanes (Paraffins)

- 9.12. Methane and Ethane
- 9.13. Physical Properties
- 9.14. Chemical Properties
- 9.15. Uses
- 9.16. Alkenes (Olefins)
- 9.17. Properties of Ethene
- 9.18. Uses of Ethylene
- 9.19. Alkynes
- 9.20. Properties of Alkynes
- 9.21. Uses of Acetylene
- 9.22. Methanol and Ethanol
- 9.23. Uses of Alcohol
- 9.24. Carboxylic Acids
- 9.25. Properties
- 9.26. Uses of Carboxylic Acids

## 9.0 INTRODUCTION

There are millions of compounds of carbon that are obtained from plant and animal kingdoms. For example, sugars, alcohol, indigo, urea, blood, albumin, penicillin, etc. In 1808, Berzelius introduced the term 'organic' to describe all the substances grouped under plant and animal kingdoms.

In eighteenth century, organic compounds were simply isolated from plant and animal sources.

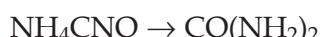
**Table 9.1 Compounds and its Sources**

Compounds	Sources
Sucrose	Beet root
Citric acid	Lemon

Lactic acid	Sour milk
Benzoic acid	Gum benzoin
Urea	Urine
Protein, Fat	Animals
Fuel gases, Naphtha, Gasoline	Petroleum
Methanol, Acetone	Wood
Dyes, Drugs, Perfumes	Coal
Sugar, Starch and Cellulose	Plants

In the earlier days, it was believed that organic compounds were produced by the influence of a mysterious force called the 'vital force'. In 1828, a German chemist named Friedrich Wohler actually synthesized the organic compound urea, in the

laboratory, from ammonium cyanate, for the first time.



This synthesis gave a death blow to the vital force theory. Many organic compounds were subsequently prepared in the laboratory. Urea is found in the urine of mammals.

### 9.1 COMPARISON BETWEEN ORGANIC AND INORGANIC COMPOUNDS

**Table 9.2 Difference between organic and inorganic compound**

Organic Compound	Inorganic Compound
1. All these compounds contain carbon which is found to combine with a limited elements like S, H <sub>2</sub> , O <sub>2</sub> , N <sub>2</sub> , halogens, etc.	Most of them do not contain carbon. They are found to have elements in the combined form.
2. There are over 3 million compounds.	There are approximately 100,000 known inorganic compounds.
3. All of them are covalent compounds.	They are ionic, covalent and co-ordinate compounds.
4. They are combustible in air.	They are generally non-combustible.
5. Organic reactions are slow as they are covalent.	Inorganic reactions are generally fast because they are ionic.
6. They are generally soluble in organic solvents like benzene, alcohol, ether; but insoluble in water.	They are generally insoluble in organic solvents, but soluble in water.
7. They have low melting point and low boiling point.	They generally have high melting point and high boiling point.
8. Homologous series of compounds are found.	Homologous series are not found.
9. Organic compounds show isomerism.	They do not generally show isomerism.
10. Molecular reactions are slow, due to the presence of linkages. These reactions never proceed to completion.	Reactions are fast if ionic and slow if covalent in nature.

11. They undergo polymerization reaction.	They do not exhibit polymerization.
12. Organic compounds have characteristic colour and odour.	Most of them are colourless and odourless.

### 9.2 APPLICATIONS OF ORGANIC CHEMISTRY

Organic compounds are applied in our day-to-day life materials such as soaps and shampoos ; medicines like aspirin, proteins, fats, vitamins, petroleum, explosives, dyes, insecticides, etc.

Organic compounds are now defined as compounds that are based on carbon. They usually contain carbon – carbon and carbon – hydrogen bonds.

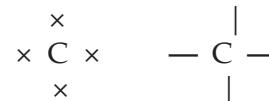
*Carbon is the basis for organic chemistry, as it occurs in all living organisms.*

### 9.3 UNIQUE NATURE OF CARBON ATOMS

There are several million organic compounds, but only about a quarter of million inorganic compounds. The reason for more organic compounds are the bonding properties of carbon.

#### Tetravalency of The Carbon Atom

Each carbon atom usually forms four covalent bonds. Thus, a carbon atom can connect to as many as four other atoms. Since it can neither lose nor gain electrons to attain octet, it forms covalent bonds by sharing its four electrons with other atoms. This tendency is known as tetravalency of carbon. Carbon can bond too many other types of atoms, including hydrogen, oxygen and nitrogen.



The Lewis structure above shows methane, the simplest organic compound. The carbon atom has four valence electrons, and it obtains four more electrons by forming four covalent bonds with the four hydrogen atoms.

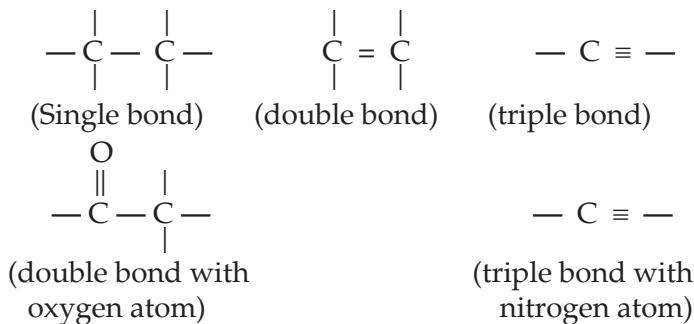
In addition to single bond, carbon can also form double and triple bonds with other carbon atoms. In the case of carbon-carbon single bond one pair of electrons are shared between two carbon atoms. In a carbon–carbon double bond, two pairs of electrons are shared between two atoms. In a triple bond, three pairs of electrons are shared between the two carbon atoms.

Carbon is a non-metal that can bond with itself and many other chemical elements, forming nearly ten million compounds.

### Catenation

The carbon atom has a unique ability to link successively to other carbon atoms to form chains of varying lengths and shapes. It is called 'catenation'. This property is responsible for the presence of large number of organic compounds.

The molecule that contains single carbon–carbon bonds is saturated. Molecules that contain double or triple carbon–carbon bonds are unsaturated. The carbon atoms on either side of the double or triple bond are bonded to less than four atoms each.



**Elemental carbon can take the form of one of the hardest substances (diamond) or one of the softest (graphite).**

Carbon's unique bonding properties allow the formation of a variety of structures, including chains and rings of many shapes and sizes. The figure above indicates the many shapes that can be formed based on carbon atoms.

### 9.4 TYPES OF ORGANIC COMPOUNDS

The organic compounds have been classified as hydrocarbons based on certain groups of atoms; the group of atoms is called functional groups. The hydroxyl group, -OH, is an example of a functional group, and compounds with this group attached to a carbon framework are called alcohols.

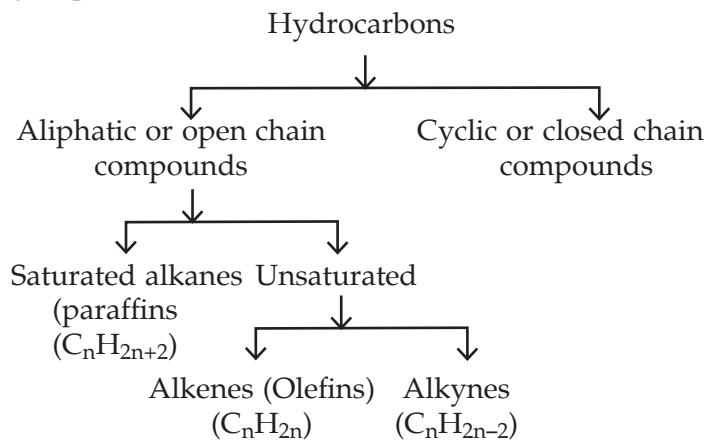
Carbon is made in the interiors of stars, though it was not produced in the Big Bang.

Some of the main functional groups are as follows:

### 9.5 CLASSIFICATION OF ORGANIC COMPOUNDS

The number of organic compounds is very large. It will be very difficult to study them under one heading. The study can be made convenient by dividing the

organic compounds into different classes or groups. Structural similarities and properties of members of a group can serve as the basis for such a classification.

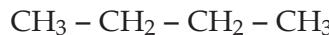
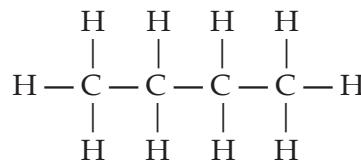


Organic compounds can be broadly classified into acyclic, aliphatic or open chain and cyclic or closed chain compounds.

### 1. Acyclic Compounds

The compounds in which all carbon atoms are arranged in a straight or branched chain are called acyclic or open chain compounds. They are also called aliphatic compounds. (Greek *aliphatos* means 'fats').

Example: butane (C<sub>4</sub>H<sub>10</sub>). The carbon chain in butane can be found in a straight chain form or in branched form.



These aliphatic compounds are further divided into saturated compounds and unsaturated compounds.

### Saturated Compounds

Good example for open chain saturated compounds are alkanes. They are represented by the formula C<sub>n</sub>H<sub>2n+2</sub>. In these compounds all the four valencies of carbon are satisfied by single covalent bonds. And it does not contain any functional group.

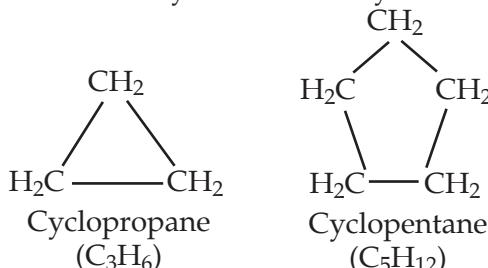
### Unsaturated Compounds

In unsaturated carbon compounds the valencies of carbon are satisfied by double and triple bonds. For example, compounds with double bond are called alkenes; and compounds with triple bond are called alkynes. The general formula for ALKENES is C<sub>n</sub>H<sub>2n</sub> and the general formula for ALKYNES is C<sub>n</sub>H<sub>2n-2</sub>.

## 2. Cyclic Compounds

**Carbon has the highest melting or sublimation point of the elements. The melting point of diamond is  $\sim 3550^{\circ}\text{C}$ , with the sublimation point of carbon around  $3800^{\circ}\text{C}$**

In some compounds, carbon atoms are linked in such a manner as to form a closed chain. Such compounds are called cyclic compounds. Based on ring structure, these cyclic compounds are further classified into homocyclic and heterocyclic compounds.



### (a) Homocyclic Compounds:

Carbon compounds have limitless uses. In its elemental form, diamond is a gemstone and used for drilling/cutting; graphite is used in pencils, as a lubricant, and to protect against rust; while charcoal is used to remove toxins, tastes and odours. The isotope Carbon-14 is used in radiocarbon dating.

In these compounds, the ring structure is made up of only carbon atoms. It is further divided into:

#### (i) Alicyclic compounds

A hydrocarbon with a cyclic chain of carbon atoms

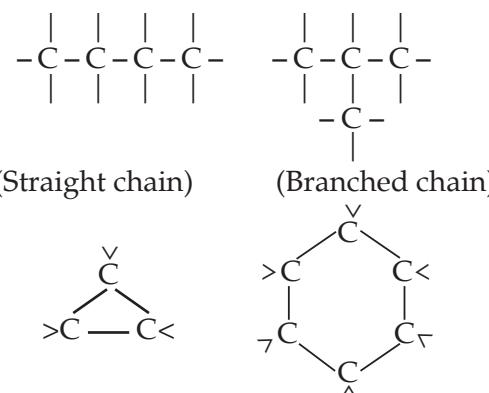
with single bonds between all the atoms is called alicyclic compound. They have the general formula  $\text{C}_n\text{H}_{2n}$ . For example, cyclopropane.

#### (ii) Aromatic compounds.

A hydrocarbon with a cyclic chain of carbon atoms with alternate single and double bonds is called aromatic compound. For example, benzene.

#### (b) Heterocyclic Compounds :

Cyclic compounds in which the ring structure is made up of atoms like nitrogen, oxygen and sulphur, in addition to carbon atoms, are called heterocyclic compounds.



#### Alkyl Group:

An alkyl group is obtained by removing one atom of hydrogen from an alkane molecule.

An alkyl group is named by replacing the suffix -ane of the alkane with the suffix -yl

Name	Molecular formula	Abbreviated formula	Structural formula
Methyl	$-\text{CH}_3$	$-\text{CH}_3$	$\begin{array}{c} \text{H} \\   \\ -\text{C}-\text{H} \\   \\ \text{H} \end{array}$
Ethyl	$-\text{CH}_2\text{CH}_3$	$-\text{CH}_2-\text{CH}_3$	$\begin{array}{cc} \text{H} & \text{H} \\   &   \\ -\text{C} & -\text{C}-\text{H} \\   &   \\ \text{H} & \text{H} \end{array}$
Propyl	$-\text{CH}_2\text{CH}_2\text{CH}_3$	$-\text{CH}_2\text{CH}_2-\text{CH}_3$	$\begin{array}{ccc} \text{H} & \text{H} & \text{H} \\   &   &   \\ -\text{C} & -\text{C} & -\text{C}-\text{H} \\   &   &   \\ \text{H} & \text{H} & \text{H} \end{array}$
Butyl	$-\text{CH}_2-\text{CH}_2-\text{CH}_2-\text{CH}_3$	$-\text{CH}_2-\text{CH}_2-\text{CH}_2-\text{CH}_3$	$\begin{array}{cccc} \text{H} & \text{H} & \text{H} & \text{H} \\   &   &   &   \\ -\text{C} & -\text{C} & -\text{C} & -\text{C}-\text{H} \\   &   &   &   \\ \text{H} & \text{H} & \text{H} & \text{H} \end{array}$

Fig. 9.3 : Structural Formula of Alkyl Groups

## Functional Group

A functional group is an atom or a group of atoms that defines the structure (or the properties of a particular family) of organic compounds.

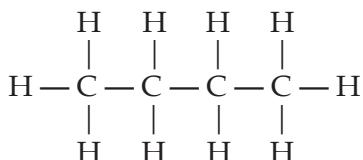
### Structure

Hydrocarbons are generally explained by their structural formula.

The formula that shows how atoms of different elements are linked together in a molecule is known as structural formula.

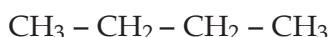
**For example :**

### Structural Formula



### Condensed Formula

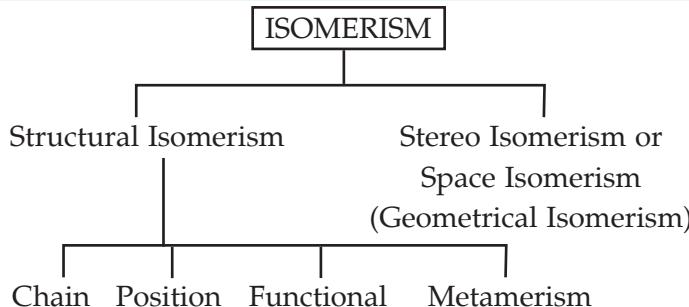
The bigger structural formula is condensed or abbreviated as follows



### Skeleton Structure

Skeleton structure represents the linkage between the carbon atoms C – C – C – C

## 9.7 ISOMERS



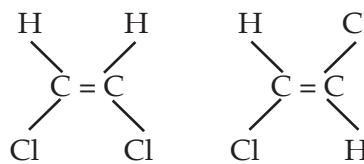
Isomers are molecules with the same number and kinds of atoms but with different arrangements. (iso= same, meros = made of the same parts)

Structural isomers have the same molecular formula but different structural formulas.

In one formula, the two carbons are connected to one another by a single covalent bond; in the other formula, each carbon is connected to the oxygen. When we complete the valency, the arrangements of atoms produce two different substances ethanol and methoxymethane.

Ethanol and methoxymethane are structural isomers. They have the same molecular formula but different structural formula. Ethanol and methoxymethane were different in their physical and chemical properties as well.

Difference in the arrangement of atoms or groups in space

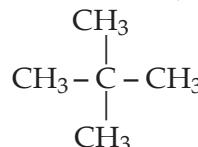
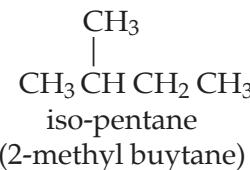
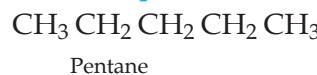


## 9.6 DIFFERENT TYPES OF STRUCTURAL ISOMERISM

### 1. Chain Isomerism

Chain isomers are compounds that have the same molecular formula but have different carbon skeletons.

**For Example :** Pentane  $\text{C}_5\text{H}_{12}$



### 2. Position Isomerism

Position isomers are compounds that have the same formula, carbon skeleton and functional groups but have the functional groups located at different positions along the carbon skeleton.

### 3. Functional Isomerism

Functional isomers are compounds that have the same molecular formula but have different functional groups.

### 4. Metamerism

It arises due to the unequal distribution of alkyl groups on either side of the functional groups in the molecules.

## 9.7 HOMOLOGOUS SERIES

A group or class of organic compounds related to each other by a general molecular formula contributes homologous series.

An ethylene group ( $\text{CH}_2$ ) is added to the hydrocarbon results in the formation of higher number of carbon atoms.

Similarly, by removing  $\text{CH}_2$  group from a hydrocarbon results in the formation of lesser number of carbon atoms.

## 9.8 CHARACTERISTICS OF HOMOLOGOUS SERIES

Homologous series have the following characteristics :

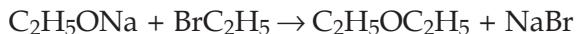
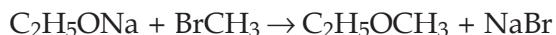
1. All members of a series contain same elements and the same functional groups.
2. All the members of a homologous series can be represented by a general formula.

Homologous series	Formula
Alkanes	$C_nH_{2n+2}$
Alkenes	$C_nH_{2n}$
Alkynes	$C_nH_{2n-2}$

Fig. 9.4 : Homologous series

3. All the members of a homologous series can be prepared by similar methods.

Ethers are prepared by alkyl halides.



4. All the members of a homologous series usually undergo similar chemical reactions.
5. Successive members in a series differ by a  $-CH_2$  group.
6. The physical properties of the member of a homologous series vary in a regular way down the series. For example, boiling point, melting point, and density of the alkane series vary in a regular way with increasing number of carbon atoms.

Examples of homologous series:

### 1. Alkane

Table 9.5 Alkane series and its molecular formula

Alkane Group	Molecular formula ( $C_nH_{2n+2}$ )
Methane	$CH_4$
Ethane	$C_2H_6$
Propane	$C_3H_8$
Butane	$C_4H_{10}$
Pentane	$C_5H_{12}$
Hexane	$C_6H_{14}$
Heptane	$C_7H_{16}$
Octane	$C_8H_{18}$
Nonane	$C_9H_{20}$
Decane	$C_{10}H_{22}$

### 2. Alkene

Table 9.6 Alkene series and its molecular formula

Alkene Groups	Molecular Formula ( $C_nH_{2n}$ )
Ethene	$C_2H_4$
Propene	$C_3H_6$
Butene	$C_4H_8$
Pentene	$C_5H_{10}$
Hexene	$C_6H_{12}$
Heptene	$C_7H_{14}$
Octane	$C_8H_{16}$
Nonane	$C_9H_{18}$
Decane	$C_{10}H_{20}$

### 3. Alkyne

Table 9.7 Alkyne series and its molecular formula

Alkyne Group	Molecular Formula ( $C_nH_{2n-2}$ )
Ethyne	$C_2H_2$
Propyne	$C_3H_4$
Butyne	$C_4H_6$
Pentyne	$C_5H_8$
Hexyne	$C_6H_{10}$
Heptyne	$C_7H_{12}$
Octyne	$C_8H_{14}$

### 4. Alcohols

Table 9.8 Alcohol group and its molecular formula

Alcohol Groups	Molecular Formula ( $C_nH_{2n+1}OH$ )
Methanol	$CH_3OH$
Ethanol	$C_2H_5OH$
Propanol	$C_3H_7OH$
Butanol	$C_4H_9OH$
Pentanol	$C_5H_{11}OH$
Hexanol	$C_6H_{13}OH$
Heptanol	$C_7H_{15}OH$
Octanol	$C_8H_{17}OH$

## 9.9 NOMENCLATURE

In the earlier days, the system of naming of organic compounds was based on the following aspects:

- The source from which the compound was obtained.
- The structure of the compound.
- The properties of the compounds.
- The names, thus assigned, based on these aspects are called trivial or common names.



### Example:

Urea (from urine source), glucose (sweet potato), CH<sub>4</sub> (marshy place), citric acid (citrus fruit)

A more reasonable approach based on rational principles for naming compounds was proposed in 1892. This system was called the Geneva nomenclature.

➤ This system was revised twice to the new system, which is now widely accepted. This new system follows the rules published by the international union of pure and applied chemistry and is known as the IUPAC System of nomenclature.

The IUPAC name of a compound essentially consists of 3 parts :

- (a) Root words
- (b) Suffix (saturation or unsaturation and functional group)
- (c) Prefix (substituents like chloro, bromo etc.)

### Root Words

The root words indicate the number of carbons in the longest continuous carbon chain in the molecule.

**Table 9.9 Root words**

Number of carbon atoms	Root word (Greek name)
One carbon atom C1	Meth
Two carbon atoms C2	Eth
Three carbon atoms C3	Prop
Four carbon atoms C4	But
Five carbon atoms C5	Pent
Six carbon atoms C6	Hex
Seven carbon atoms C7	Hept
Eight carbon atoms C8	Oct
Nine carbon atoms C9	Non
Ten carbon atoms C10	Dec

### Suffix

The suffix indicates which functional group is present and the nature of the bond in a carbon whether it is single, double or triple bond.

**Table 9.10 Functional group and its suffix**

Nature of bond	Suffix	General name	General formula
Single bond (C-C)	-ane	alkane	C <sub>n</sub> H <sub>2n+2</sub>
Double bond (C=C)	-ene	alkene	C <sub>n</sub> H <sub>2n</sub>

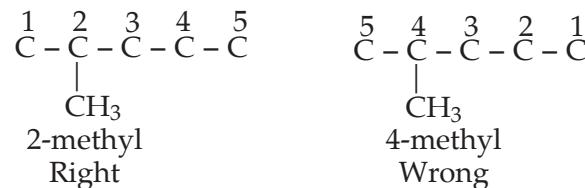
Triple bond (C≡C)	-yne	alkyne	C <sub>n</sub> H <sub>2n-2</sub>
Group (R -)	-yl	alkyl	C <sub>n</sub> H <sub>2n+1</sub>

### Prefix

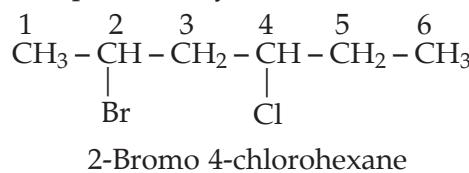
The prefix tells us the identity, location, and number of substituents attached to the carbon chain. And it also indicates whether the substituent or functional group is di, tri, tetra, etc.

1. The carbon atoms of the longest chain are numbered in both the way so that the alkyl group can attain the minimum possible number.

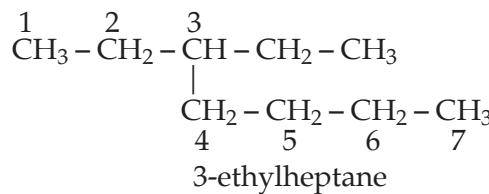
### For Example :



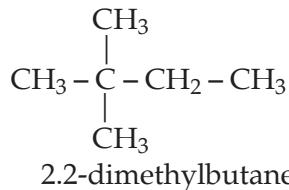
2. In case, different types of substituents are attached to the chain, they are arranged and named alphabetically.



3. The position and name of the alkyl groups is added before the parent hydrocarbon so that the alkyl groups are easily identified.

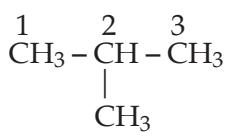


4. The centre carbon atom contains four carbon atoms, so the root name is butane and since there are two methyl branches at C-2 of the centre carbon atom, so the prefix is 2,2 - dimethyl butane.

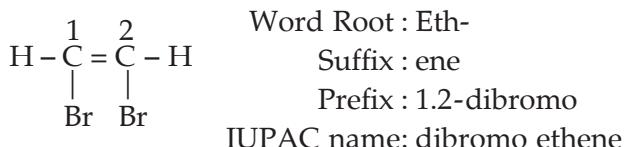
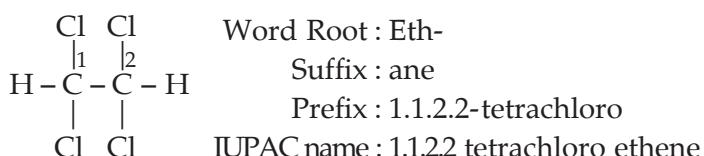
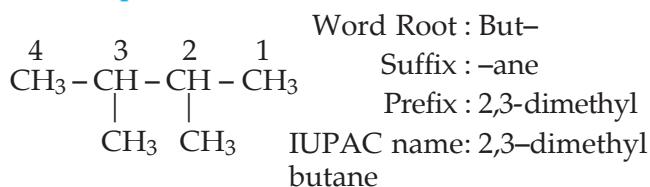
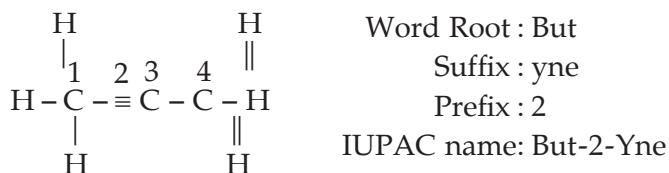
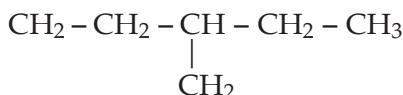


2,2-dimethylbutane

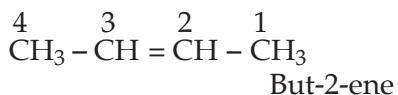
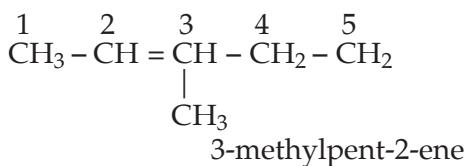
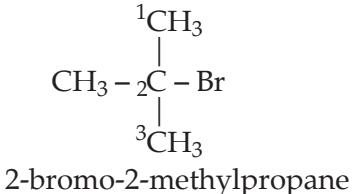
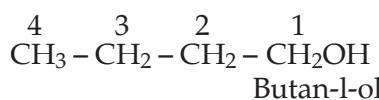
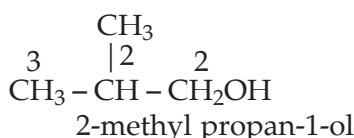
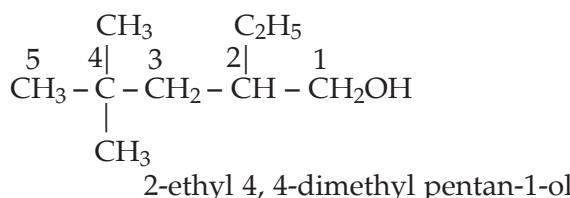
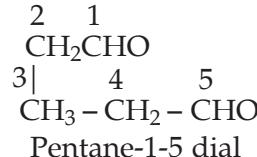
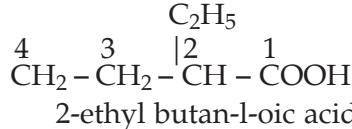
### Example : 1



Word Root : Prop-  
Suffix : ane  
Prefix : 2 methyl  
IUPAC name: 2-methyl propane

**Example : 2****Example : 3****Example : 4****Example : 5****Example : 6**

The longest chain is of 5 carbon atoms so root word is pent. The nature of carbon chain is C-C single bond, suffix – ane and the substituent is an alkyl or methyl - CH<sub>3</sub> group. Thus, name of compound is 3-methyl pentane.

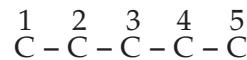
**Example : 10****Example : 11****Example : 12****Example : 13****Example : 14****Example : 15****Example : 16****Example : 17****Writing Structural Formula From IUPAC Name**

- Write the number of carbon atoms based on the carbon skeleton.
- Number the carbon atoms from any extreme end.
- Based on *ane*, *ene* or *yne* suffix the position of the bond is identified in the parent chain.
- Attach the substituent or functional group at the mentioned carbon atom.
- Satisfy the four valencies of carbon atom by attaching hydrogen atoms.
- Example:  
2 – bromo – 4 – methyl pent – 2 – ene

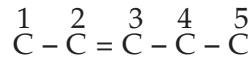
**Step 1 :** Write the carbon atom



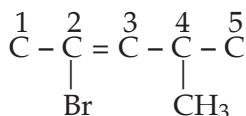
**Step 2 :** Number the carbon atom



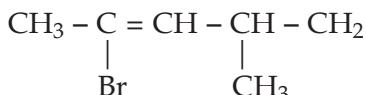
**Step 3 :** Locate the suffix



**Step 4 :** Attach substituents



**Step 5 :** Satisfy valencies by adding hydrogen atom



## 9.10 HYDROCARBONS

These compounds contain the elements carbon and hydrogen only. These compounds are divided into alkanes, alkenes and alkynes.

### 9.11 ALKANES (PARAFFINS)

Alkanes are hydrocarbons, which contain only carbon and hydrogen. These compounds are made up of carbon-carbon single bond. They have a general molecular formula  $\text{C}_n\text{H}_{2n+2}$ , where 'n' represents the number of carbon atoms in the molecule. The root words of first four numbers are taken from the common name and for higher member the root word is derived from Greek or Latin number indicating the total number of carbon atoms. For alkanes, the IUPAC names are the root word with the suffix '-ane'.

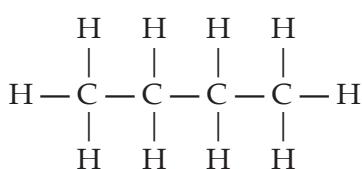
#### Source of Alkane

The two most important natural sources of alkanes are petroleum and natural gas. Petroleum is a complex liquid mixture of organic compounds, many of which are alkanes or cycloalkanes.

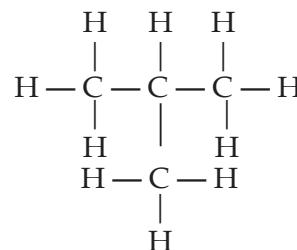
Natural gas, often found associated with petroleum deposits, consists mainly of methane (about 80%) and ethane (5% to 10%), with lesser amounts of some higher alkanes. Propane is the major constituent of liquefied petroleum gas (LPG), a domestic fuel used in homes. Butane is another choice for domestic use.

#### Isomerism in Alkanes, Single Bond Hydrocarbons

#### Isomerism of butane



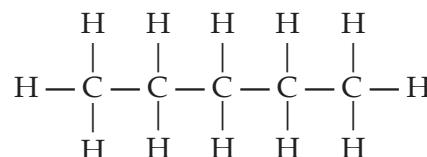
n – butane ( $\text{C}_4\text{H}_{10}$ ) , where 'n' stands for normal and IUPAC name is BUTANE



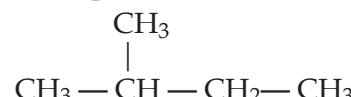
Iso – butane ( $\text{C}_4\text{H}_{10}$ ) IUPAC name – 2-methyl propane

#### Isomers of Pentane ( $\text{C}_5\text{H}_{12}$ )

Pentane can form three isomers by having five carbon atoms arranged in three different ways.

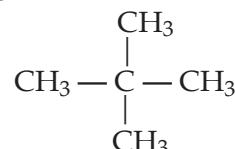


IUPAC name is pentane



IUPAC Name 2-methyl butane

Iso pentane refers to a centre carbon attached to a least of 3 carbons



IUPAC Name 2,2 dimethyl propane

Neo at least one carbon atom is attached to four other carbon atoms

## 9.12 METHANE AND ETHANE

#### Occurrence of Methane And Ethane

They are commonly found in nature wherever bacteria decompose organic matter in the absence of oxygen, as in marshes, swamps, or the muddy sediment of lakes. Hence, it has the common name marsh gas.

It is similarly formed from bacteria in the digestive tracts of certain ruminant animals, such as cows. The scale of methane production by bacteria is considerable. The earth's atmosphere contains an average of 1 part per million of methane. Because our planet is small and because methane is light compared to most other air constituents ( $\text{O}_2$ ,  $\text{N}_2$ ) one would expect most of the methane to have escaped from our atmosphere.

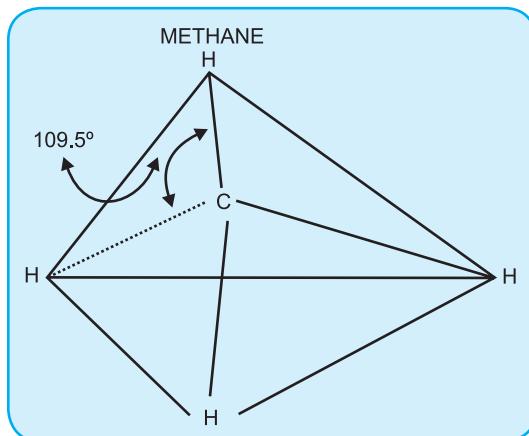
The amount of methane in the atmosphere has reached much higher levels, up to several parts per

million in recent years. The peak concentrations come in the early morning and late afternoon, directly correlated with the peak of automobile traffic.

### Structure of Methane

An atom surrounded by four groups places these four groups at the corners of a tetrahedron, giving bond angles of approximately  $109.5^\circ$ .

The carbon atom in methane ( $\text{CH}_4$ ) is bonded to four hydrogen atoms, pointing to the corners of a tetrahedron.

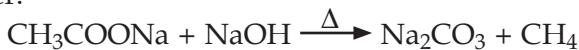


To draw the three-dimensional tetrahedron on a two-dimensional page, place two of the bonds in the plane of the page, one bond in front, and one bond behind. Then, use the drawing conventions first presented.

A solid line is used for bonds in the plane, a wedge is used for a bond in front of the plane, and a dashed line is used for a bond behind the plane.

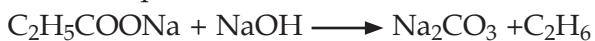
### Laboratory Preparation of Methane

Methane is prepared by heating the mixture of anhydrous sodium acetate and soda lime (a mixture of sodium hydroxide and calcium oxide) in a hard glass tube. It is collected by the downward displacement of water.



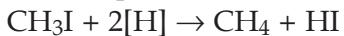
### Laboratory Preparation of Ethane

Ethane is prepared by heating the mixture of sodium propionate and soda lime in a hard glass tube using a Bunsen burner. The gas is collected by the downward displacement of water.

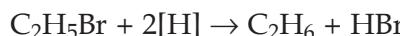


### Other Methods of Preparation of Methane and Ethane

When iodomethane is reduced with nascent hydrogen, methane is produced.



Similarly, when bromomethane is reduced with nascent hydrogen, ethane is produced.



When water is added to aluminium carbide at room temperature, methane is produced.



### Wurtz Reaction

When methyl iodide or methyl bromide is warmed with sodium metal in the presence of dry ether, ethane is produced.



## 9.13 PHYSICAL PROPERTIES

### Methane

- (a) It is a colourless and odourless gas.
- (b) It is insoluble in water, soluble in alcohol and ether.
- (c) It is lighter than air.
- (d) Its melting point is  $-183^\circ\text{C}$  and its boiling point is  $-162^\circ\text{C}$ .
- (e) It is combustible but not a supporter of combustion.

### Ethane

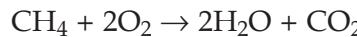
- (a) It is a colourless, odourless and non-poisonous gas.
- (b) Its boiling point is  $-89^\circ\text{C}$  and its melting point is  $-172^\circ\text{C}$ .
- (c) It is sparingly soluble in water but wholly soluble in organic solvents like alcohols, acetone and ether.

## 9.14 CHEMICAL PROPERTIES

### Methane

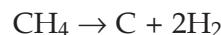
#### Combustion

It burns in air with a blue flame producing carbon dioxide gas.



#### Pyrolysis

When heated above  $1000^\circ\text{C}$ , methane gets decomposed into carbon and hydrogen. When pyrolysis occurs in alkanes, the process is termed as cracking.



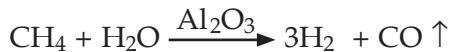
When it is heated to  $1500^\circ\text{C}$  in the absence of air along with a catalyst, it is broken down into a lower alkane.





## Reaction with Steam

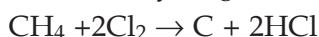
When steam is passed in the presence of heated aluminium oxide, methane liberates hydrogen gas, which is used for the synthesis of ammonia. This reaction results in the evolution of carbon monoxide.



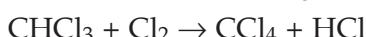
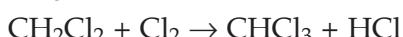
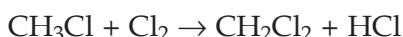
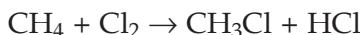
## Reaction with Chlorine

Depending on the intensity of sunlight, methane reacts in two ways with chlorine.

- (a) In bright sunlight, chlorine reacts with methane to give carbon and hydrogen chloride.



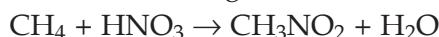
- (b) In the presence of diffused sunlight, methane undergoes substitution reaction with chlorine, forming four products namely, methyl chloride, methylene chloride, chloroform and carbon tetrachloride. It is a chain reaction.



The substitution products can be separated by fractional distillation because of the appreciable difference in their respective boiling points.

## Nitration

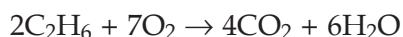
In vapour phase, at  $400^{\circ}\text{C}$ , methane reacts with concentrated nitric acid to give nitromethane.



## Ethane

### Combustion

It burns with oxygen with a blue flame producing carbon dioxide gas.



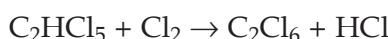
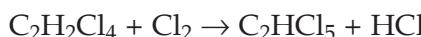
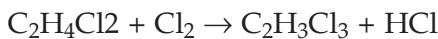
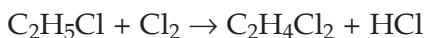
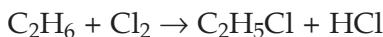
## Pyrolysis

On heating ethane under  $500^{\circ}\text{C}$  temperature or in the presence of a catalyst  $\text{Al}_2\text{O}_3$  in absence of air it is broken down into lower member.



## Reaction with chlorine

Ethane also reacts with chlorine in the same way as methane, producing successively chloroethane, dichloroethane, trichloroethane, pentachloroethane and hexachloroethane.



## NOTE

All alkanes react with all the halogens in a similar manner, producing the corresponding substituted products.

## 9.15 USES

### Methane

- (a) As a fuel.
- (b) As a source of hydrogen for the manufacture of ammonia. In the preparation of carbon black used in printing ink and in the manufacture of tyres.

### Ethane

- (a) As a good fuel.
- (b) It is used to prepare ethene, ethanol, ethanol and ethanoic acid.
- (c) It forms ethyl chloride, which is used to make tetraethyllead.

## 9.16 ALKENES (OLEFINS)

Alkenes are hydrocarbons with carbon – carbon double bond. They have a general molecular formula  $\text{C}_n\text{H}_{2n}$ . Where 'n' represents the number of carbon atoms in the molecule. Their IUPAC names are the root word with a suffix – 'ene'.

### Isomers In Alkene

Alkenes having more than 4 carbon atoms can form isomers.

For example: butene has three isomers:

1.  $\text{CH}_3\text{CH}_2\text{CH}=\text{CH}_2$
2.  $\text{CH}_3\text{CH}=\text{CHCH}_3$
3.  $\begin{array}{c} \text{CH}_2 = \text{C} - \text{CH}_3 \\ | \\ \text{CH}_3 \end{array}$

There are two types of isomerism shown by alkenes :

- (i) Chain isomerism :



1 - Butene

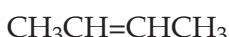


2 – methyl propene

(ii) Position isomerism :



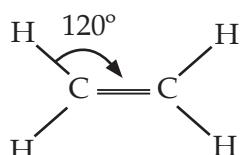
1 - butene



2 - butene

### Ethene (Ethylene) $\text{C}_2\text{H}_4$

Molecular formula  $\text{C}_2\text{H}_4$ , molecular weight 28, is the first member of the alkene family. It is used for the preservation and ripening of fruits like apples, oranges, bananas and mangoes.

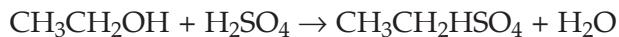
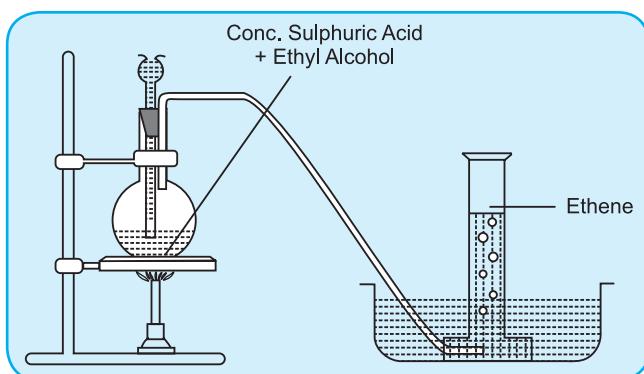


### Structure of Ethene

### Preparation Of Ethene (Ethylene)

(i) By dehydration of ethyl alcohol

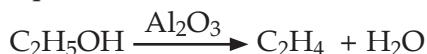
Ethylene is prepared by heating ethyl alcohol with excess of concentrated sulphuric acid at a temperature of  $160^\circ\text{C}$  to  $170^\circ\text{C}$  in a round bottom flask provided with a thistle funnel and a delivery tube. The gas coming out of the delivery tube is passed through a wash bottle containing sodium or potassium hydroxide solution. Pure ethylene is collected by the downward displacement of water.



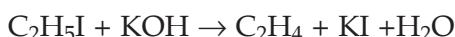
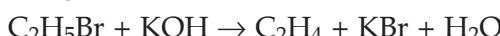
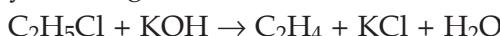
(ii) By dehydration

### Industrial Preparation :

Ethylene is prepared by passing the ethanol vapours through a tube containing alumina ( $\text{Al}_2\text{O}_3$ ) at  $300^\circ\text{C}$  temperature.



(iii) Dehydrohalogenation:



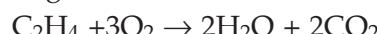
### 9.17 PROPERTIES OF ETHENE

#### Physical Properties

- It is a colourless, odourless gas.
- It produces anaesthetic effects on inhalation.
- It is insoluble in water and soluble in alcohol and ether.
- It is lighter than air.
- It has a boiling point of  $120^\circ\text{C}$  and melting point is  $-169^\circ\text{C}$ .
- It is combustible but not a supporter of combustion.

#### Chemical Properties

1. **Combustion** : It burns in air with a luminous flame producing water and carbon dioxide.



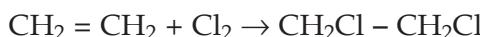
2. **Addition Reaction** : Being unsaturated, ethylene undergoes addition reaction. During this process, the double bond changes to single bond through ionic mechanism. It shows the following addition reaction.

(a) **Addition of hydrogen**: When hydrogen is added to ethylene, in the presence of nickel or platinum or palladium as catalyst at  $200^\circ\text{C} - 300^\circ\text{C}$ , ethane is formed.

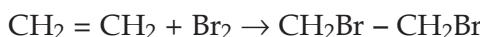


(b) **Addition of halogen**: Chlorine and bromine react differently with ethylene.

(i) When chlorine is added to ethylene, ethylene dichloride is formed.



(ii) When bromine water is added, ethylene decolourises forming ethylene dibromide. This is used as a test to identify ethylene.



(c) **Addition of hydrogen halide,  $\text{HX}$**

(i) When hydrogen chloride is added, ethylene forms ethyl chloride.



(ii) When hydrogen bromide is added, ethylene forms ethyl bromide.



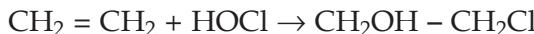
(d) Addition of concentrated sulphuric acid,  $\text{H}_2\text{SO}_4$ .

When concentrated sulphuric acid is added, ethylene gives ethyl hydrogen sulphate.



(e) Addition of hypochlorous acid,  $\text{HOCl}$

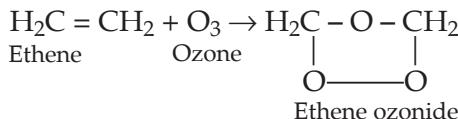
When hypochlorous acid is added, ethylene gives ethylene chlorohydrin.



**3. Oxidation :** Ethylene is readily oxidized by cold dilute alkaline potassium permanganate (Baeyer's reagent) to give ethylene glycol. This reaction is another identification test for ethylene.

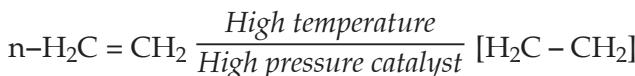


**4. Ozonolysis:** When ozone is dissolved in a solution of ethene in ether, ethene ozonide is formed.



**5. Polymerisation :** When two or more molecules of the same compound associate to form a bigger molecule, the reaction is called polymerisation.

Ethene polymerises to produce polyethene.



**Table 9.11 Molecular formula and its IUPAC names**

Molecular formula	Condensed formula	Structural formula	Trivial name (Common name)	IUPAC name
$\text{C}_2\text{H}_4$	$\text{CH}_2 = \text{CH}_2$	$  \begin{array}{c}  \text{H} \quad \text{H} \\    \quad   \\  \text{H} - \text{C} = \text{C} - \text{H}  \end{array}  $	Ethylene	Ethene
$\text{C}_3\text{H}_6$	$\text{CH}_3 - \text{CH} = \text{CH}_2$	$  \begin{array}{c}  \text{H} \quad \text{H} \quad \text{H} \\    \quad   \quad   \\  \text{H} - \text{C} - \text{C} = \text{C} - \text{H} \\    \\  \text{H}  \end{array}  $	Propylene	Propene
$\text{C}_4\text{H}_8$	$\text{CH}_3 - \text{CH}_2 - \text{CH} = \text{CH}_2$	$  \begin{array}{c}  \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\    \quad   \quad   \quad   \\  \text{H} - \text{C} - \text{C} - \text{C} = \text{C} - \text{H} \\    \quad   \\  \text{H} \quad \text{H}  \end{array}  $	Butylene	Butene

### Homologous Series of Alkenes

#### Ethyne

Acetylene is the first member of the alkyne series. Acetylene is prepared by the action of water on

### 9.18 USES OF ETHYLENE

1. Ethylene is used as a fuel.
2. It is used for preservation and ripening of fruits like apples, oranges, bananas and mangoes.
3. It is used in the manufacture of polythene.
4. It is used for the manufacture of ethanol, glycol etc.
5. It is used in making epoxyethane (used in the manufacture of detergents).
6. It is used for producing oxy-ethylene flame, which is used for cutting and welding of metals.
7. It is used as anesthetic (i.e., a mixture of ethylene and oxygen).

### 9.19 ALKYNES

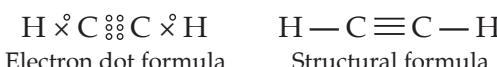
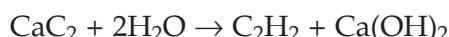
Alkynes are hydrocarbons with carbon – carbon triple bond. They have the root word with the suffix –yne'. It comes with the general formula  $\text{C}_n\text{H}_{2n-2}$

Alkynes are hydrocarbons obtained from the natural gas.

Alkynes with four or more than four carbon atoms can form isomers. They show both position as well as chain isomerism.

calcium carbide. Lumps of  $\text{CaC}_2$  are placed in a flask fitted with a thistle funnel and a delivery tube. Water is dropped from the funnel and the acetylene evolved is passed through acidified copper sulphate solution

to remove the impurities. It is then collected by the downward displacement of water.



The shape of ethyne molecule is linear

### PROPERTIES OF ETHYNE

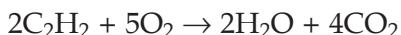
#### Physical Properties

- It is a colorless gas with a sweet odour.
- It is lighter than air.
- It liquefies at  $-84^{\circ}\text{C}$ .
- Its boiling point is  $75^{\circ}\text{C}$ .
- It is sparingly soluble in water but soluble in alcohol and ether.
- It is combustible but not a supporter of combustion.

#### Chemical Properties

##### 1. Combustion

It burns in air with a lot of smoke, producing carbon dioxide.



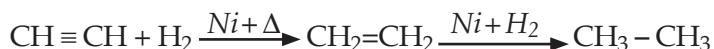
##### 2. Addition Reactions

Since acetylene contains a triple bond, it is a highly unsaturated compound. Two molecules of hydrogen or chlorine or bromine or hydrogen chloride or hydrogen bromide can be added. Two addition products are formed.

During the addition process, the triple bond changes to double bond and then to single bond.

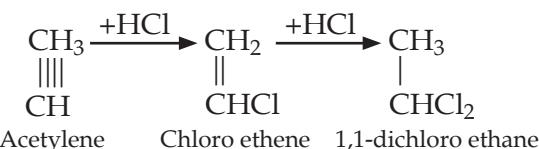
###### (a) Addition of hydrogen:

At  $250^{\circ}\text{C}$ , when hydrogen is added in the presence of nickel as a catalyst, ethene and then ethane is formed.



###### (b) Addition of halogens :

- (i) With chlorine, acetylene first changes to acetylene dichloride and with further addition of chlorine, acetylene tetrachloride is formed.



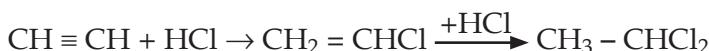
Acetylene reacts vigorously with chlorine gas in the presence of sunlight to give out flames.

(ii) Acetylene decolourises bromine water and acetylene dibromide is formed. On further addition of bromine acetylene tetrabromide is formed.

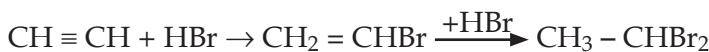


(c) Addition of hydrogen halides,  $\text{HX}$ :

- (i) When hydrogen chloride is added, acetylene first forms vinyl chloride and on further addition of hydrogen chloride, ethylidene chloride is formed.



- (ii) With hydrogen bromide, acetylene first gives vinyl bromide and on further addition of hydrogen bromide, ethylidene bromide is formed.



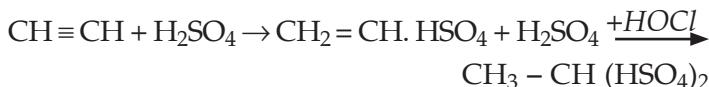
(d) Addition of hypochlorous acid,  $\text{HOCl}$

With hypochlorous acid, acetylene first forms acetylene chlorohydrins and then gives dichloroacetaldehyde.



(e) Addition of sulphuric acid :

With sulphuric acid, acetylene first forms vinyl hydrogen sulphate and then gives ethylidene hydrogen sulphate.



### 9.21 USES OF ACETYLENE

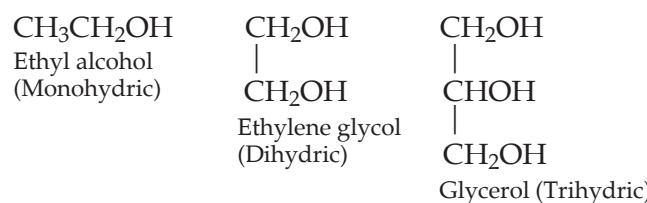
1. It is used as anesthetic. It is used in oxyacetylene blow pipe flame for welding and cutting operations. It is used as a fuel.
2. Used for artificial ripening of fruits for the preparation of acetaldehyde, acetic acid, and alcohol etc.
3. In the manufacture of plastics like PVC, synthetic rubber, synthetic fibre (Orlon).

#### Alcohols

Alcohols are organic compounds with general formula  $\text{R} - \text{OH}$ , where  $\text{R}$  is an alkyl group. These may be considered as hydroxyl derivatives of hydrocarbons in which one or more hydrogen atoms are replaced by hydroxyl (-OH) groups.

Alcohols are classified according to the number of hydroxyl groups present. The alcohols containing

one hydroxyl group are called monohydric alcohol, the alcohols containing two hydroxyl groups are known as dihydric alcohols, and the alcohols containing three hydroxyl groups are known as trihydric alcohols.



**Table 9.12 Homologous Series of Alcohols**

Common Name	Molecular formula	Abbreviated formula	Structural formula	IUPAC formula
Methyl alcohol	$\text{CH}_3\text{OH}$	$\text{CH}_3 - \text{OH}$	$\begin{array}{c} \text{H} \\   \\ \text{H} - \text{C} - \text{O} - \text{H} \\   \\ \text{H} \end{array}$	Methanol
Ethyl alcohol	$\text{C}_2\text{H}_5\text{OH}$	$\text{CH}_3 - \text{CH}_2 - \text{OH}$	$\begin{array}{c} \text{H} \quad \text{H} \\   \quad   \\ \text{H} - \text{C} - \text{C} - \text{O} - \text{H} \\   \quad   \\ \text{H} \quad \text{H} \end{array}$	Ethanol
Propyl alcohol	$\text{C}_3\text{H}_7\text{OH}$	$\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{OH}$	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\   \quad   \quad   \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{O} - \text{H} \\   \quad   \quad   \\ \text{H} \quad \text{H} \quad \text{H} \end{array}$	Propanol
Butyl alcohol	$\text{C}_4\text{H}_9\text{OH}$	$\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{OH}$	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\   \quad   \quad   \quad   \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{O} - \text{H} \\   \quad   \quad   \quad   \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$	Butanol

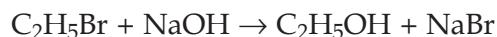
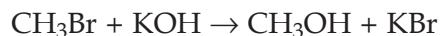
## 9.22 METHANOL AND ETHANOL

### Occurrence

Alcohols are obtained by synthesis, for example, it can be obtained by the destructive distillation of wood and from the fermentation process of sugar (obtained from the molasses).

### Laboratory Preparation

Alcohols can be prepared by the hydrolysis of alkyl halides.



Methanol (Large Scale Method)

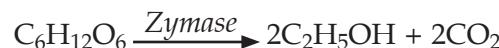
Ethanol

Hydration of Ethene

### Fermentation of Carbohydrates

In the fermentation process the enzyme invertase present in yeast hydrolyses sucrose into glucose and fructose, the enzyme zymase present in yeast

converts glucose and fructose to ethyl alcohol and carbon dioxide. This process of preparation of ethanol is called fermentation.



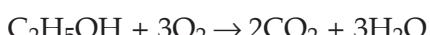
### Properties of Alcohol

#### Physical Properties

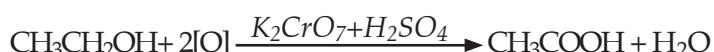
- It is a colourless liquid having pleasant smell.
- It is miscible with water in all proportions.
- It is non-conductor of electricity because it does not contain ions.
- Ethanol is lighter than water.
- If methyl alcohol is consumed, it causes blindness and even death.
- Consuming Ethyl alcohol damages the liver and kidney and affects parts of the brain, which control our muscular movements. It also gives temporary relief from tiredness.

### Chemical Properties

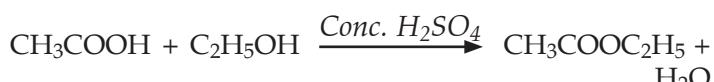
**1. Combustion :** Ethanol burns in air with a blue flame to form carbon dioxide and water.



**2. Oxidation :** Ethanol is oxidised to ethanoic acid by an acidic solution of potassium dichromate.

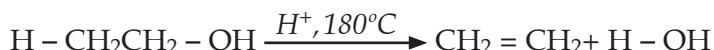


**3. Reaction With Carboxylic Acids :** When ethanol is warmed with ethanoic acid in presence of a few drops of concentrated sulphuric acid, sweet smell of ethyl ethanoate (ethyl acetate) is produced.



The process of the formation of an ester (ethyl ethanoate) by the combination of ethanol with ethanoic acid is known as esterification.

**4. Dehydration of Alcohols :** Alcohols can be dehydrated by heating them with a strong acid. For example, when ethanol is heated at  $180^\circ\text{C}$  with a small amount of concentrated sulphuric acid, a good yield of ethylene is obtained.



### Denatured Alcohol or Methylated Spirit

Rectified spirit is made unfit for drinking purpose, by adding 5% methyl alcohol, 0.5% pyridine and some colouring matter and is called denatured alcohol or methylated spirit. It is used for the preparation of paints and varnishes as a solvent.

### Methanol

It was earlier prepared from wood by distillation process and hence is called as wood alcohol. At present, however, methanol is manufactured from carbon monoxide and hydrogen.

### Spurious Alcohol

This is a type of liquor prepared by improper distillation. It contains large proportions of methanol in a mixture of alcohols, and it is used as the solvent for the paint and varnishes.

### Absolute Alcohol

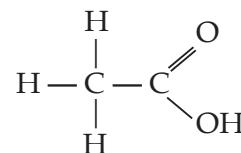
Commercial alcohol is a constant-boiling mixture containing 95% ethanol and 5% water and cannot be further purified by distillation. To remove the remaining water to obtain Absolute Alcohol, one adds quick lime ( $\text{CaO}$ ), which reacts with water to form calcium hydroxide but does not react with ethanol.

### 9.23. USES OF ALCOHOL

- It is used as a solvent for fats and many other organic compounds.
- In the preparation of esters, used as perfumes.
- In the manufacture of chemicals such as chloroform, Chloral, Iodoform, Ether, Acetic acid, ethylene etc.
- In the form of fuel as power alcohol as beverage.
- As preservative for biological specimens and as antiseptic in hospitals.
- As an antifreeze for automobile radiators.
- As a solvent for pharmaceutical preparations and a constituent of medicines.
- As a fluid for scientific apparatus, e.g., thermometers, spirit levels etc., in which alcohol colored with a dye is used

### 9.24 CARBOXYLIC ACIDS

Organic compounds which contain the carboxylic functional group ( $-\text{COOH}$ ) are called the carboxylic acids. The general formula is



The name *carboxyl* is derived from carbonyl ( $\text{C}=\text{O}$ ) and hydroxyl ( $-\text{OH}$ ) groups because in the carboxyl group these two groups are directly bonded to each other. The properties of the carboxyl group are not simply as those of carbonyl and hydroxyl groups combined. The two groups interact to give carboxylic acids, which show their own distinctive properties.

Table 9.13 Formulae and nomenclature of acids

Formula	Common name	IUPAC name
$\begin{array}{c} \text{O} \\    \\ \text{H} - \text{C} - \text{OH} \end{array}$	Formic acid	Methanoic acid
$\begin{array}{c} \text{O} \\    \\ \text{CH}_3 - \text{C} - \text{OH} \end{array}$	Acetic acid	Ethanoic acid
$\begin{array}{c} \text{O} \\    \\ \text{CH}_3\text{CH}_2 - \text{C} - \text{OH} \end{array}$	Propionic acid	Propanoic acid
$\begin{array}{c} \text{O} \\    \\ \text{CH}_3\text{CH}_2\text{CH}_2 - \text{C} - \text{OH} \end{array}$	Butyric acid	Butanoic acid

Carboxylic acids are classified as monocarboxylic acids, dicarboxylic acids, tricarboxylic acids etc. according to the number of  $-COOH$  groups present in the molecule 1, 2, 3, etc. respectively. The long chain monocarboxylic acids are commonly called fatty acids because many of them are obtained by the hydrolysis of animal fats or vegetable oils. For example, Stearic acid, Palmitic acid.

### Nomenclature of Mono-carboxylic Acids

The simple carboxylic acids are better known by their common names. The common names are usually derived from the Latin or Greek words that indicate the original source of the acid. For example:

Formic acid is present in ants (Latin: Formica = ants) and Acetic acid is present in vinegar (acetum = vinegar).

Butyric acid was named from the fact that butter fat (Latin: butyric = butter) is a significant source of this acid.

### IUPAC System

1. The longest carbon chain containing the carboxylic group is considered as the parent hydrocarbon.
2. The names of the carboxylic acids are derived from the names of the parent hydrocarbons by replacing the terminal 'e' by 'oic acid'.

Name of carboxylic acid = Name of corresponding alkane - e + oic acid

**Example :** (1) HCOOH

Methane + Oic acid  $\rightarrow$  Methanoic Acid

(2) CH<sub>3</sub>COOH

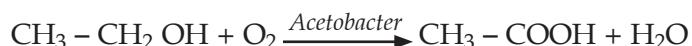
Ethane + Oic acid  $\rightarrow$  Ethanoic Acid

## 9.25 PROPERTIES

### Preparation

#### Quick Vinegar Process

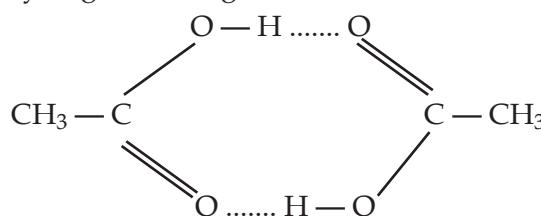
A dilute solution of acetic acid (3 – 7%) is known as vinegar. It is prepared by the fermentation of ethyl alcohol by the bacteria acetobacter in presence of air.



### Physical Properties

- (i) Acetic acid is a colorless liquid (boiling point 118°C) with a sharp 'vinegar odour' and sour taste.
- (ii) It is miscible with water, ethyl alcohol, and ether in all proportions.
- (iii) When cooled sufficiently, it forms 'ice like' crystals, which melt at 16.7°C. hence the pure anhydrous acid is usually called glacial acetic acid (glacial = ice).
- (iv) It is a good solvent for sulphur, phosphorus, iodine and many organic compounds.

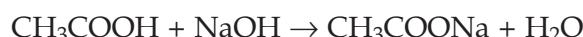
- (v) Acetic acid exists as a dimer due to inter-molecular hydrogen bonding between two molecules.



### Chemical Properties

#### 1. Salt Formation

Acetic acid reacts with alkalies to form corresponding salts. Acetic acid can donate a proton and form salts with bases. This shows the acidic nature of acetic acid.



#### 2. Formation of Acid Halides

Acetic acid reacts with phosphorus pentahalide to form acid halides (or with thionyl chloride, SOCl<sub>2</sub>, to form acetyl chlorides).



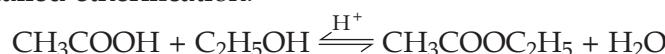
#### 3. Formation of Amides

Acetic acid reacts with ammonia to give salts, which on heating yield acetamide.



#### 4. Formation of Esters

Acetic acid reacts with ethyl alcohol in the presence of a strong acid catalyst like H<sub>2</sub>SO<sub>4</sub> to form esters. The reaction is reversible and the forward reaction is called etherification.



#### 5. Decarboxylation :

Decarboxylation is the elimination of CO<sub>2</sub> from a carboxylic acid.

When sodium salts of acetic acid are heated with soda lime (NaOH + CaO) alkanes are formed.



### Tests for Carboxylic Acid

1. When Ethanoic acid is warmed with ethanol in the presence of a few drops of concentrated sulphuric acid, a sweet smelling ester called ethyl ethanoate or ethylacetate is formed.
2. Acetic acid produces wine red colour when a neutral solution of ferric chloride is added to it.

## 9.26 USES OF CARBOXYLIC ACID

1. Used as a coagulant for rubber latex.
2. In the manufacture of plastics, rayon, drugs and silk.

3. Dilute aqueous solution (3 – 7 %) of ethanoic acid. is called vinegar, which is used to preserve food.
4. Pure ethanoic acid is used as a solvent and chemical reagent.
5. As cellulose ethanoate, it is used in making photographic films and rayon.

6. It is widely used in the manufacture of textiles.
7. It is used in the preparation of white lead and Paris green, a bright blue-green pigment used mainly as a wood preservative.

### INTEXT QUESTIONS

1. How organic compounds were obtained in the eighteenth century?
2. Write few sources of organic compounds.
3. How are cyclic compounds classified?
4. What is meant by tetravalency and catenation?
5. What are hydrocarbons?
6. Give at least one example for the following:  
Single bond compound, double bond compound and triple bond compound.
7. Write any two differences between organic and inorganic compounds.
8. What is meant by saturated and unsaturated compounds?
9. Define isomerism.
10. Write the formula and draw the structure of benzene molecule.
11. Which part of an organic compound determines physical and chemical properties?
12. What is meant by an alkyl group?
13. What is a homologous series?
14. State the source of alkanes.
15. How does methane acts as a green house gas?
16. Give the general formula of alkanes.
17. Write the molecular formula, electron dot formula and structural formula of methane and ethane.
18. How is methane prepared in the laboratory?
19. How is ethane prepared in the laboratory?
20. What is substitution reaction?
21. Write the uses of methane.
22. Write the uses of ethane.
23. Write few physical properties of alkane.
24. Convert methane into chloroform.
25. Write the molecular formula and electron dot formula and structural formula of ethane.
26. Give a balanced equation for the lab preparation of ethylene. How is the gas collected?
27. Give the conditions and the main product formed by hydrogenation of ethylene.
28. How is ethanol converted into ethene using a solid dehydrating agent?
29. Write few physical properties of ethene.
30. How would you convert ethene into 1, 2 – dibromoethane?
31. Write the balanced equation for ethene combines with hydrogen chloride.
32. Write the source of alkyne.
33. How is acetylene prepared in the laboratory.
34. Classify the following compounds as alkanes, alkenes and alkynes.
35. Give a chemical test to distinguish between saturated and unsaturated compounds.
36. What are the products formed when ethyne is reacted with chlorine and bromine.
37. Name the hydrocarbon which is a tetrahedral molecule.
38. What are alcohols? State their sources.
39. Give the general formula of monohydric alcohol.
40. Give the electron dot, abbreviated formula and structure of second member of the alcohol group.
41. State the method of preparation of ethanol.
42. How is ethanol prepared by fermentation process?
43. What is the effect of ethanol on human body?
44. How is absolute alcohol obtained?
45. Name the products formed and give appropriate chemical equations for the following :  
(a) Sodium reacting with ethyl alcohol.  
(b) Ethanol oxidized by acidified potassium dichromate.
46. What are carboxylic acids?
47. Write the common name, IUPAC name and formula of one monocarboxylic acid and one dicarboxylic acid.
48. Give the structural formulae and IUPAC name of acetic acid. What is glacial acetic acid?
49. How is acetic acid prepared from ethanol and acetylene?



### POINTS TO REMEMBER

- ◆ Alkanes have the general formula  $C_nH_{2n+2}$ .
- ◆ Alkenes have the general formula  $C_nH_{2n}$ .
- ◆ Alkanes have the general formula  $C_nH_{2n-2}$ .
- ◆ Alkanes are also called as parafins.



- ◆ Alkenes are called as olefins.
- ◆ Methane is also called marsh gas.
- ◆ Carboxylic acids are used as a coagulant for rubber latex.

- ◆ Mixture of ethylene and oxygen is used as anesthetic.
- ◆ Alkene reacts with ozone to form ozonide.

## EXERCISES

### A. FILL IN THE BLANKS

1. Two types of isomerism are ..... and .....
2. General formula of Alkanes .....
3. Alkanes are also called as .....
4. Methane is also called .....
5. The end product of nitration of methane are ..... and water.
6. Conversion of ethene to ethane is an example of .....

### B. MULTIPLE CHOICE TYPE QUESTIONS

Choose the correct answer from the options given below :

1. Paraffins is another name of
 

(a) Alkanes	(b) Alkynes
(c) Alkenes	(d) Alcohols
2. Ethylene is used as :
 

(a) Anesthetic	(b) Antibiotic
(c) Ore	(d) Oxidising agent
3. Alkenes are also called as :
 

(a) Paraffins	(b) Cyclo paraffins
(c) Olefins	(d) None of these
4. Chemical formula of ethyl alcohol
 

(a) $\text{CH}_3\text{OH}$	(b) $\text{C}_4\text{H}_9\text{OH}$
(b) $\text{C}_2\text{H}_5\text{OH}$	(d) $\text{HCHO}$
5. IUPAC name of  $\text{HCOOH}$  is :
 

(a) Methanoic acid	(b) Acetic acid
(c) Ethanoic acid	(d) Carbonic acid

### C. CHOOSE THE MOST APPROPRIATE ANSWER FOR EACH OF THE FOLLOWING

- (i) Identify the statement which does **not** describe the property of alkenes :
  - (A) They are unsaturated hydrocarbons.
  - (B) They decolourise bromine water.
  - (C) They can undergo addition as well as substitution reactions.
  - (D) They undergo combustion with oxygen forming carbon dioxide and water. [2015]
- (ii) Explain the bonding in methane molecule using electron dot structure. [2015]

### D. MATCH THE COLUMNS

Column I	Column II
1. Alkanes	$\text{C}_n\text{H}_{2n+2}$
2. Alkenes	Marsh gas
3. Alkynes	$\text{C}_n\text{H}_{2n-2}$
4. Methane	Anesthetic
5. Ethylene	$\text{C}_n\text{H}_{2n}$

### E. GIVE THE BALANCED CHEMICAL EQUATION FOR THE FOLLOWING

Give balanced chemical equations for the following conversions :

- (i) Ethanoic acid to Ethyl ethanoate.
- (ii) Calcium carbide to Ethyne
- (iii) Sodium ethanoate to Methane. [2015]

### F. VERY SHORT QUESTIONS

1. Write the general formula for a saturated hydrocarbon and give one example of a saturated hydrocarbon with its structural formula. [ICSE 2004]
2. Name a compound, which will give acetylene gas when treated with water. [ICSE 2004]
3. Ethane, Ethene, Ethanoic acid, Ethyne, Ethanol [ICSE 2005]
  - (a) The compound with – OH as the part of its structure.
  - (b) The compound with – COOH as the part of its structure.
  - (c) Homologue of homologous series with general formula  $\text{C}_n\text{H}_{2n}$ .
4. State one use of acetylene. [ICSE 2007]
5. Using appropriate catalysts, ethane can be oxidised to an alcohol, an aldehyde and an acid. Name the alcohol, aldehyde and acid formed when ethane is oxidised. [ICSE 2008]
6. Why is pure acetic acid known as glacial acetic acid? [ICSE 2008]
7. What type of compound is formed by the reaction between acetic acid and an alcohol? [ICSE 2008]
8. Name the organic compound prepared by each of the following reactions : [ICSE 2012]

- (a)  $\text{CH}_3\text{COONa} + \text{NaOH} \xrightarrow{\text{CaO}}$   
 (b)  $\text{CaC}_2 + \text{H}_2\text{O} \longrightarrow$   
 (c)  $\text{C}_2\text{H}_5\text{Br} + \text{KOH}(\text{alc}) \longrightarrow$   
 (d)  $\text{C}_2\text{H}_5\text{Br} + \text{KOH}(\text{aq}) \longrightarrow$   
 (e)  $\text{C}_2\text{H}_5\text{OH} + \text{CH}_3\text{COOH} \xrightarrow[\text{H}_2\text{SO}_4]{\text{Conc.}}$
9. Name the following : [ICSE 2012]  
 (a) Second member of alkene series  
 (b) First member of alkane series  
 (c) Third member of aldehyde series  
 (d) Second member of carboxylic acid.  
 (e) Fourth member of alcohol series.
10. Draw different isomers having the following molecular formula: [ICSE 2012]  
 I.  $\text{C}_5\text{H}_{12}$  (chain)  
 II.  $\text{C}_4\text{H}_8$  (position)
11. What is denatured alcohol?
12. Give two important uses of ethanol.
13. Write equations for : [ICSE 2012]  
 (i) Preparation of ethanol by hydration of  $\text{C}_2\text{H}_4$   
 (ii) Preparation of acetic acid from ethanol
14. Using their structural formulae identify the functional group by circling them :  
 (i) Dimethyl ether      (ii) Propanone [2015]
15. Name the following :  
 (i) Process by which ethane is obtained from ethene.  
 (ii) A hydrocarbon which contributes towards the **greenhouse** effect.  
 (iii) Different reactions that takes place when ethanol is treated with acetic acid.  
 (iv) The property of elements by virtue of which atoms of the element can link to each other in the form of a long chain or ring structure.  
 (v) Reaction when an alkyl halide is treated with alcoholic potassium hydroxide.

### G. SHORT QUESTIONS

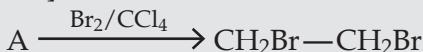
1. Draw the structural formula of ethane and vinegar.
2. Write balanced equation when ethane is burnt in air. [ICSE 2004]
3. Write the equation for the preparation of ethylene from the ethyl alcohol. [ICSE 2004]
4. Draw the structural formula of a compound with two carbon atoms in each of the following cases.  
 [ICSE 2005]  
 (a) An alkane with a carbon-to-carbon single bond.

- (b) An alcohol containing two carbon atoms.  
 (c) An unsaturated hydrocarbon with a carbon to carbon triple bond.
5. Write the equation for the following laboratory preparations: [ICSE 2005]  
 (a) Ethane from sodium propionate  
 (b) Ethene from iodoethane  
 (c) Ethyne from calcium carbide  
 (d) Methanol from iodomethane
6. Write the equation for the preparation of carbon tetrachloride from methane. [ICSE 2007]
7. Draw the structural formula of ethyne. [ICSE 2007]
8. Draw the structural formulae of the two isomers of butane. Give the correct IUPAC name of each isomer. [ICSE 2007]
9. Name the organic compound prepared by each of the following reactions : [ICSE 2008]  
 (i)  $\text{C}_2\text{H}_5\text{COONa} + \text{NaOH} \longrightarrow$   
 (ii)  $\text{CH}_3\text{I} + 2\text{H} \longrightarrow$   
 (iii)  $\text{C}_2\text{H}_5\text{Br} + \text{KOH} \longrightarrow$   
 (iv)  $\text{CO} + 2\text{H}_2 \longrightarrow$   
 (v)  $\text{CaC}_2 + 2\text{H}_2\text{O} \longrightarrow$
10. Write the equations for the following reactions : [ICSE 2008]  
 (i) Calcium carbide and water.  
 (ii) Ethene and water(steam).  
 (iii) Bromoethane and an aqueous solution of sodium hydroxide.
11. Addition reactions and substitution reactions are types of organic reactions. Which type of reaction is shown by : [ICSE 2008]  
 (i) Ethane      (ii) Ethene ?
12. Write the equation of the complete combustion of ethane. [ICSE 2008]
13. Give the chemical equation for : [ICSE 2009]  
 (a) The laboratory preparation of methane from sodium acetate.  
 (b) The industrial preparation of methanol from water gas.  
 (c) The reaction of one mole of ethene with one mole of chlorine gas.  
 (d) The preparation of ethyne from 1,2 – dibromoethane.
14. (a) Define isomerism [ICSE 2009]  
 (b) Give the IUPAC name of the isomer  $\text{C}_4\text{H}_{10}$  that has a branched chain. [ICSE 2009]
15. Draw the structural formula for each of the following : [ICSE 2010]  
 (a) Ethanoic acid  
 (b) But – 2 – yne



## H. LONG QUESTIONS

- How is the structure of alkynes different from that of alkenes? [ICSE 2007]
- Distinguish between the saturated hydrocarbon ethane and the unsaturated hydrocarbon ethene by drawing their structural formulae.[ICSE 2008]
- State how the following conversions can be carried out: [ICSE 2009]
  - Ethyl chloride to Ethyl alcohol
  - Ethyl chloride to Ethene
  - Ethene to ethyl Alcohol
  - Ethyl alcohol to Ethene
- Compound A is bubbled through bromine dissolved in carbon tetrachloride and the product is  $\text{CH}_2\text{Br} - \text{CH}_2\text{Br}$  [ICSE 2010]



- Draw the structural formula of A.
  - What type of reaction has A undergone?
  - What is your observation?
  - Name the compound formed when steam reacts with A in the presence of phosphoric acid.
  - What is the procedure for converting the product of (b) (IV) back to A?
- Write balanced chemical equations of the following : [ICSE 2011]
    - Monochloro-ethane is hydrolysed with aqueous KOH.
    - A mixture of soda lime and sodium acetate is heated.
    - Ethanol under high pressure and low temperature is treated with acidified potassium dichromate.
    - Water is added to calcium carbide.
    - Ethanol reacts with sodium at room temperature.

## RECENT BOARD QUESTIONS

### Question 1.

**(a) Choose the correct answer from the options given below :**

[2011,12,13,15]

- The unsaturated hydrocarbons undergo :
  - a substitution reaction
  - an oxidation reaction
  - an addition reaction
  - none of the above
- The number of C-H bonds in ethane molecule are :
  - Four
  - Six
  - Eight
  - Ten
- Which of the following properties do not match with elements of the halogen family ?
  - They have seven electrons in their valence shell.
  - They are highly reactive chemically.
  - They are metallic in nature.
- Identify the statement that is incorrect about alkanes :
  - They are hydrocarbons.
  - There is a single covalent bond between carbon and hydrogen.
  - They can undergo both substitution as well as addition reactions.
  - On complete combustion they produce carbon dioxide and water.

### Question 2.

**(a) Choose the correct word/phrase from within the brackets to complete the following sentences :** [2011]

- The catalyst used for conversion of ethene to ethane is commonly ..... (nickel/iron/cobalt).
- When acetaldehyde is oxidized with acidified potassium dichromate, it forms..... (ester/ethanol/acetic acid).
- Ethanoic acid reacts with ethanol in presence of concentrated  $\text{H}_2\text{SO}_4$ , so as to form a compound and water. The chemical reaction which takes place is called..... (dehydration/hydrogenation/esterification)
- Write the equation for the reaction taking place between 1, 2-dibromoethane and alcoholic potassium hydroxide.
- The product formed when ethene gas reacts with water in the presence of sulphuric acid is..... (ethanol/ethanal/ethanoic acid).

**(b) Write balanced chemical equations for the following :** [2011]

- Monochloro ethane is hydrolysed with aqueous KOH.
  - Ethanol under high pressure and low temperature is treated with acidified potassium dichromate.
  - Ethanol reacts with sodium at room temperature.
  - Preparation of ethanol from ethyl chloride.
- (c) Give the structural formula for the following :** [2012]
- Methanoic acid
  - Ethanal
  - Ethyne
  - Acetone
  - 2-methyl propane.

### Question 3.

**(a) Give balanced equations for the laboratory preparations of the following organic compounds :** [2013]

- (i) A saturated hydrocarbon from iodomethane.  
 (ii) An unsaturated hydrocarbon from an alcohol.  
 (iii) An unsaturated hydrocarbon from calcium carbide.  
 (iv) An alcohol from ethyl bromide.

**(b) Give the structural formulae for the following : [2013]**

- (i) An isomer of n-butane (ii) 2-propanol  
 (iii) Diethyl ether

**(c) Give reasons for the following : [2013]**

- (i) Methane does not undergo addition reactions, but ethene does.  
 (ii) Ethyne is more reactive than ethane.  
 (iii) Hydrocarbons are excellent fuels.

#### Question 4.

**(a) State the conditions required for the following reactions to take place : [2014]**

- (i) Catalytic hydrogenation of ethyne.  
 (ii) Preparation of ethyne from ethylene dibromide.

**(b) Give the appropriate answer for the following :**

- (i) Give the I.U.P.A.C. name of acetylene.



- (ii) Hydrocarbons containing a  $\text{—C}\text{—}$  functional group.  
 (iii) A gaseous hydrocarbon commonly used for welding purposes.

**(c) Give the Complete chemical Reaction for the following : [2014]**

- (i) Preparation of ethane from sodium propionate.  
 (ii) Preparation of ethanol from monochloroethane and aq. sodium hydroxide.

#### Question 5.

**(a) Give reasons as to why : [2011]**

- (i) Almost 90% of all known compounds are organic in nature.  
 (ii) It is dangerous to burn methane in an insufficient supply of air.

**(b) An organic compound with vapour density = 94 contains.**

$C = 12\cdot67\%$ ,  $H = 2\cdot13\%$ , and  $Br = 85\cdot11\%$ . Find the molecular formula.

[Atomic mass :  $C = 12$ ,  $H = 1$ ,  $Br = 80$ ]

They are diatomic in their molecular form.

**(c) Give the structural formula of the following : [2014]**

- (i) Ethanol (ii) 1-propanol  
 (iii) Ethanoic acid (iv) 1, 2, dichloroethane

**(d) Name the gase : [2014]**

- (i) This gas is used for welding purposes.  
 (ii) This gas is also a saturated hydrocarbon.

#### Question 6.

**(a) Give balanced chemical equation for the conversion of : [2015]**

- (i) Ethanoic acid to ethyl ethanoate.

**(b) Using their structural formulae identify the functional group by circling them : [2015]**

- (i) Dimethyl ether (ii) Propanone

**(d) Name the following : [2015]**

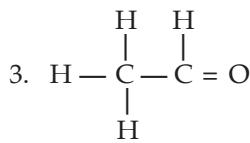
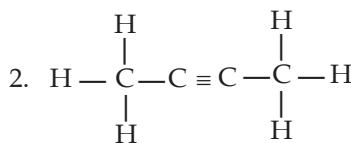
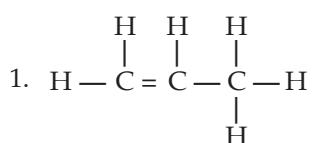
- (i) Process by which ethane is obtained from ethene.  
 (ii) A hydrocarbon which contributes towards the greenhouse effect.  
 (iii) Distinctive reaction that takes place when ethanol is treated with acetic acid.  
 (iv) The property of elements by virtue of which atoms of the element can link to each other in the form of a long chain or ring structure.  
 (v) Reaction when an alkyl halide is treated with alcoholic potassium hydroxide.  
 (vi) Explain the bonding in methane molecule using electron dot structure.

#### Question 7.

**(a) Identify the term/substance in each of the following : [2016]**

- (i) The catalyst used in the conversion of ethyne to ethane.  
 (ii) The type of reactions alkenes undergo.

**(b) (i) Write the IUPAC names of each of the following : [2016]**



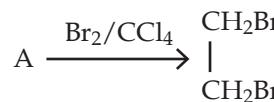
**(c) Write a balanced chemical equation for each of the following :**

- (i) Burning of ethane in plentiful supply of air.  
 (ii) Heating of ethanol at  $170^\circ\text{C}$  in the presence of conc. sulphuric acid.

**(d) Give the structural formulae of each of the following :**

- (i) 2-methyl propane (ii) Ethanoic acid  
 (iii) Butan-2-o1

**(e) Equation for the reaction when compound A is bubbled through bromine dissolved in carbon tetrachloride is as follows :**



- (i) Draw the structure of A.  
 (ii) State your observation during this reacton.





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