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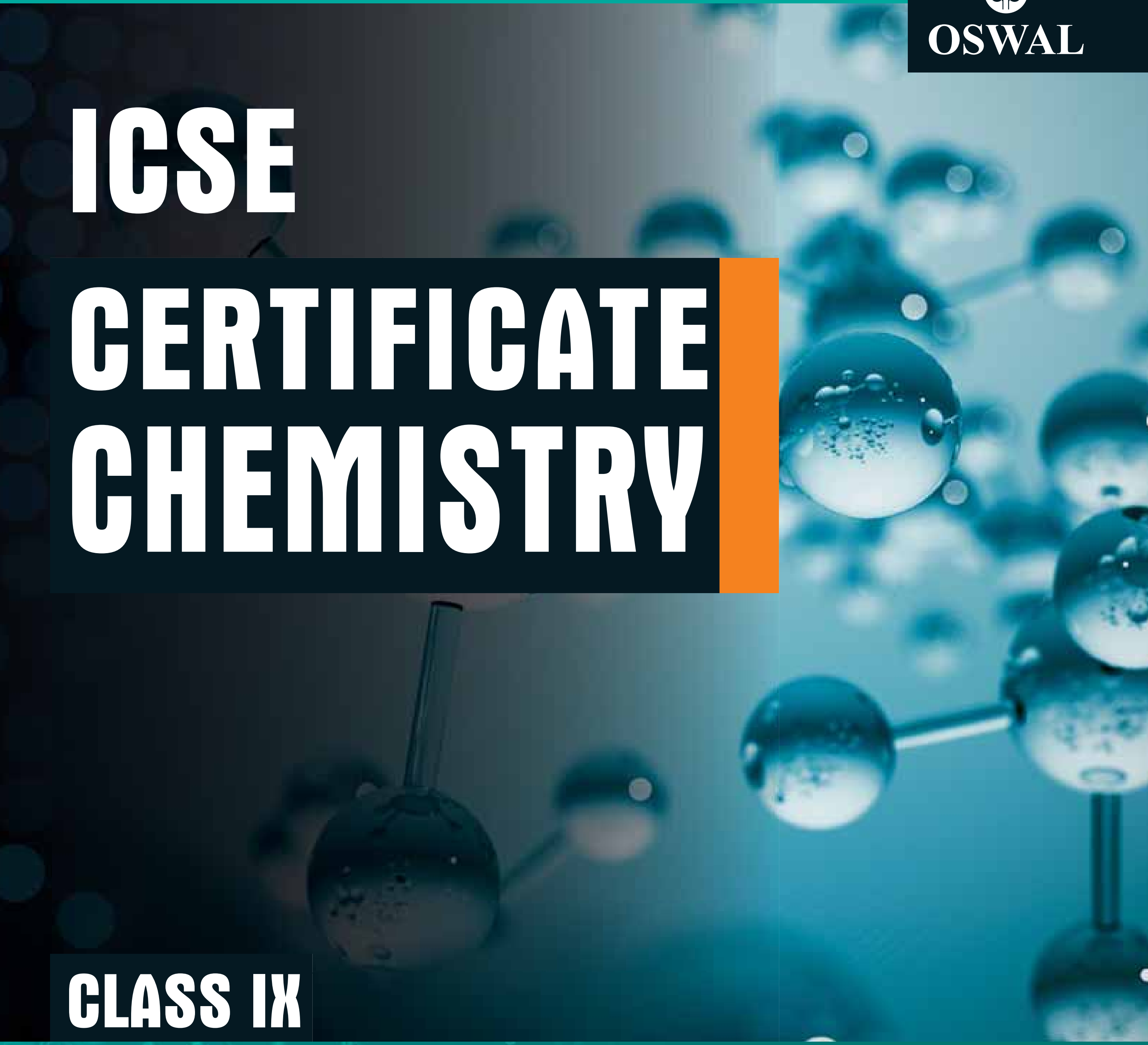

ICSE CERTIFICATE CHEMISTRY

CLASS IX

Manasa Sampat Hegde

CERTIFICATE
CHEMISTRY

In accordance with the latest syllabus prescribed by the Council for the Indian School Certificate Examinations, New Delhi.

CERTIFICATE CHEMISTRY

CLASS IX

Manasa Sampat Hegde
M. Sc., Chemistry
PGDESD

OSWAL PUBLISHERS
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Preface

Oswal publishers Certificate Chemistry for Class IX has been designed to cover the Chemistry syllabus as prescribed by the Council for the Indian School Certificate Examinations [CISCE], New Delhi. The chapters are also sequenced as per the distinct directives of the Board.

Chemistry is an important component of the Physical Science that deals with matter and its Environment. The attempt while preparing this text has been to make the subject interesting and Stimulating for young learners and inculcate a Scientific bent of mind.

The language used is simple and lucid to aid easy comprehension. Every chapter is proceeded with 'Learning outcomes' to outline the topics that students are expected to cover as per the Board directives.

Summary points after every chapter have been written to aid quick revisions. The text has been supplemented will illustrations, graphics, charts and other tools.

Self assessment Questions following previous years board paper will help students to better equip themselves to the upcoming exams and also evaluate their undertaking of the topic.

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.

SYLLABUS CLASS IX

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 reactions should be studied with reference to the reactants, products, conditions, observations and the (balanced) equations.

1. The Language of Chemistry

(i) Symbol of an element; valency; formulae of radicals and formulae of compounds. Balancing of simple chemical equations.

- *Symbol — definition; symbols of the elements used often.*
- *Valency — definition; hydrogen combination and number of valence electrons of the metals and non-metals; mono, di, tri and tetra valent elements.*
- *Radicals — definition; formulae and valencies*
- *Compounds — name and formulae.*
- *Chemical equation — definition and examples of chemicals equations with one reactant and two or three products, two reactants and one product, two reactants and two products and two reactants and three or four products; balancing of equations. (by hit and trial method).*

(ii) Relative Atomic Masses (atomic weights) and Relative Molecular Masses (molecular weights) : either – standard H atom or $1/12^{th}$ of carbon 12 atom.

- *Definitions*
- *Calculation of Relative Molecular Mass and percentage composition of a compound.*

2. Chemical changes and reactions

(i) Types of chemical changes.

- *Direct combination*
- *Decomposition*
- *Displacement;*
- *Double decomposition*

(The above to be taught with suitable chemical equations as examples).

(ii) Energy changes in a chemical change.

Exothermic and endothermic reactions with examples – evolution/absorption of heat, light and electricity.

3. Water

(i) Water as a universal solvent.
- *Solutions as 'mixtures' of solids in water; saturated solutions.*
- *Qualitative effect of temperature on solubility (e.g. solutions of calcium sulphate, potassium nitrate and sodium chloride in water).*

(ii) Hydrated and anhydrous substances.
  (a) Hydrated substances :
  Water of Crystallisation – meaning and examples
  (b) Anhydrous substances :
  Meaning and examples only
  (c) Properties :
  - *Efflorescence*
  - *Deliquescence*
  - *Hygroscopy*
  - *Removal of hardness*
    (i) By boiling
    (ii) By addition of washing soda

(Definition and examples of each of the above).

(iii) Drying and Dehydrating Agents
Meaning and examples only.

(iv) Soft Water and Hard water
- *Meaning, (in terms of action of soap)*
- *Advantages and disadvantages of soft water and hard water.*
- *Types and causes of hardness.*

4. Atomic Structure and Chemical bonding

(i) Structure of an Atom, Mass number and atomic number, Isotopes and Octet Rule.
- *Definition of an atom*
- *Constituents of an atom – nucleus (protons, neutrons) with associated electrons; mass number, atomic number.*
- *Electron distribution in the orbits – $2n^2$ rule, Octet rule. Reason for chemical activity of an atom.*
- *Definition and examples of isotopes (hydrogen, carbon, chlorine).*

(ii) Electrovalent and covalent boding, structures of various compounds – orbit structure
  (a) Electrovalent Bond
  - *Definition*
  - *Atomic orbit structure for the formation of Electrovalent compounds (e.g. NaCl, $MgCl_2$, CaO)*

  (b) Covalent Bond
  - *Definition*
  - *Atomic orbit structure for the formation of Covalent molecules on the basis of duplet and octet of electrons (examples : hydrogen, chlorine, oxygen, nitrogen, carbon tetrachloride, methane.)*

5. The Periodic Table

Dobereiner's Traids, Newland's law of Octaves, Mendeleev's contributions; Modern Periodic Law, the Modern Periodic Table (Groups and periods)
- *General idea of Dobereiner's triads, Newland's law of Octaves, Mendeleev's periodic law.*
- *Discovery of Atomic Number and its use as a basis for Modern periodic law.*

- *Modern Periodic Table (Groups 1 to 8 and periods 1 to 7).*
- *Special reference to Alkali metal (Group 1), Alkaline Earth metals (Group 2) Halogens (Group 17) and Zero Group (Group 18).*

6. Study of the First Element – Hydrogen

Position of the non-metal (Hydrogen) in the periodic table and general group characteristics with reference to valency electrons, burning, ion formation applied to the above mentioned element.

(i) Hydrogen form water, dilute acids and alkalis :

(a) Hydrogen from water :

- *The action of cold water on sodium potassium and calcium.*
- *The action of hot water on magnesium.*
- *The action of steam on aluminium zinc, and iron; (reversibility of reaction between iron and steam).*
- *The action of steam on non-metal (carbon).*

Students can be shown the action of sodium and calcium on water in the laboratory. They must be asked to make observations and write equations for the above reactions.

Application of activity series for the above mentioned reactions.

(b) Displacement of hydrogen from dilute acids :

The action of dilute sulphuric acid or hydrochloric acid on metals : Mg, Al, Zn and Fe

(To understand reasons for the using other metals and dilute nitric acid)

(c) Displacement of hydrogen from alkalis :

The action of Alkalis (NaOH, KOH) on Al, Zn and Pb – unique nature of these elements.

(ii) The preparation and collection of hydrogen by a standard laboratory method other than electrolysis.

In the laboratory preparation, the reason for suing zinc, the impurities in the gas, their removal and the precautions in the collection of the gas must be mentioned.

(iii) Industrial manufacture of hydrogen by Bosch process :

- *Main reactions and conditions.*
- *Separation of CO2 and Co from hydrogen.*

(iv) Oxidation and reduction reactions

Differences in terms of addition and removal of oxygen / hydrogen.

7. Study of Gas Laws

(i) The behaviour of gases under changes of temperature and pressure; explanation in terms of molecular motion (particles, atoms, molecules); Boyle's Law and Charles' Law; absolute zero; gas equation; simple relevant calculations.

- *The behaviour of gases under changes of temperature and pressure; explanation in terms of molecular motion (particles, atoms, molecules).*
- *Boyle's Law : statement, mathematical form, simple calculations.*
- *Charles' Law : statement, mathematical form, simple calculations.*
- *Absolute zero Kelvin scale of temperature.*
- *Gas equation* $P_1 V_1 / T_1 = P_2 V_2 / T_2$ *; simple relevant calculations based on gas equation.*

(ii) Relationship between Kelvin scale and Celsius Scale of temperature; Standard temperature and pressure.

Conversion of temperature from Celsius Scale to Kelvin scale and vice versa. Standard temperature and pressure. (Simple calcualtions).

8. Atmospheric pollution

(a) Acid rain – composition, causes and its impact.

Sulphur in fossil fuels giving oxides of sulphur when burnt. High temperatures in furnaces and internal combustion engines produce oxides of nitrogen. (Equations to be included). Acid rain affects soil chemistry and water bodies.

(b) Global warming :

Greenhouse gases – their sources and ways of reducing their presence in the atmosphere.

(Water vapour, carbon dioxide, methane and oxides of nitrogen)

(c) Ozone depletion

- *Formation of ozone – relevant equations*
- *Function in the atmosphere.*
- *Destruction of the ozone layer–chemicals responsible for this to be named but reactions not required.*

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 listed below.

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

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

Simple experiments on :

1. Action of heat on the following compounds :
   (a) copper carbonate, zinc carbonate
   (b) washing soda, copper sulphate crystals
   (c) zinc nitrate, copper nitrate, lead nitrate
   (d) ammonium chloride, iodine, ammonium dichromate

   Make observations, identify the product and make deductions where possible.
2. Action of dilute sulphuric acid on the following substances. (warm if necessary)
   (a) a metal
   (b) a carbonate
   (c) a sulphide
   (d) a sulphite

   Make observations, identify the gas evolved and make deductions
3. Apply the flame test to identify the metal in the unknown substances.
   (a) a sodium salt
   (b) a potassium salt
   (c) a calcium compound
4. Simple experiments based on hard water and soft water – identification of hardness – simple softening – by heating the temporary hard water, using washing soda and advantage of using detergents over soap in hard water.
5. Find out the sources of pollution of water bodies in the locality. Suggest preventive steps to control it.

CONTENTS

CERTIFICATE
CHEMISTRY

CHAPTER 1

LANGUAGE OF CHEMISTRY

LEARNING OUTCOMES

1.1. Introduction
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1.3. Chemical Formula
1.4. Valency
1.5. Radical
1.6. Writing Chemical Formula
1.7. Chemical Reaction and Equation
1.8. Writing a Chemical Equation
1.9. Balancing of Chemical Equations

1.1. INTRODUCTION

Chemistry is the branch of science which deals with the study of matter, investigation of their properties and the types of reaction they undergo to form new products.

Many scientists have carried out various research and put forward several theories and postulates to explain the concept of matter. Matter is defined as anything that has mass and occupies space, which are made up of atoms, the smallest particles. Further, scientists coined the term **element** for those substances which are composed of only one kind of atom. *An element is a simple and pure form of matter which cannot be decomposed into simpler substances*. A group of same types of atoms bond together to form a **molecule** of an element. The number of atoms bonded in a molecule of an element is called its **atomicity**. Molecules of elements can be classified by its atomicity as follows–

(i) Monoatomic molecule–Elements having only one atom in their molecule, *e.g.* All noble gases–Helium (He), Neon (Ne), Argon (Ar), etc. are monoatomic molecules.

(ii) Diatomic molecule–Elements having two atoms in their molecule, *e.g.* Hydrogen (H_2), Nitrogen (N_2), Oxygen (O_2), Chlorine (Cl_2), etc.

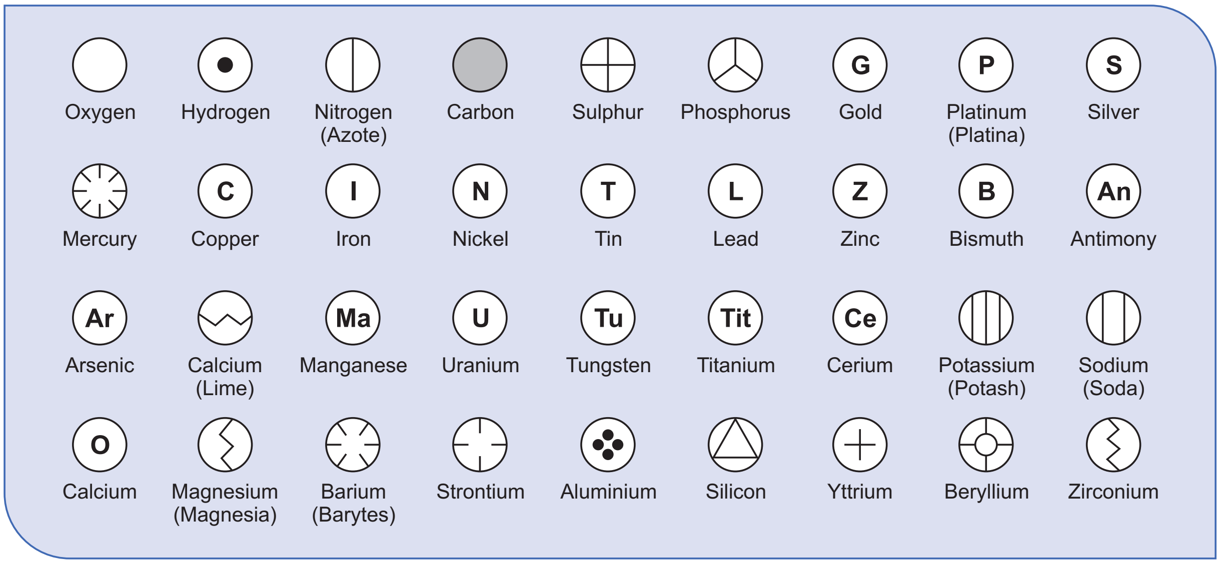
(iii) Triatomic molecule–Elements having three atoms in their molecule, *e.g.* Ozone (O_3).

(iv) Tetratomic molecule–Elements having four atoms in their molecule, *e.g.* Phosphorus (P_4).

(v) Polyatomic molecule–Elements having more than four atoms in their molecule, *e.g.* Sulphur (S_8), Boron (B_{12}), Fullerene (C_{60}), etc.

Atoms of different elements combines together in a definite proportion to form molecules of compound, *e.g.* Water (H_2O), Carbondioxide (CO_2), Calcium chloride ($CaCl_2$), etc.

The symbols of elements were first proposed by the scientist **John Dalton**. He used a circle for Oxygen atom, a plus sign in the circle for sulphur, etc. Symbols of some elements as proposed by John Dalton are shown below.



Further **Johann Berzelius** suggested that symbols should be made from one or two letters of the name of the element. For example, letter 'O' for Oxygen, 'C' for Carbon, 'H' for Hydrogen, 'Ca' for calcium, etc. But names of some elements were in different languages like Greek, Latin and German based on the place where they were discovered or their physical properties (texture or colour, etc.). Nowadays **IUPAC (International Union of Pure and Applied Chemistry)** system of chemical symbol is followed universally. Symbols of some elements are just one alphabet while for some elements two alphabets are used. The first alphabet is always written as capital while the second alphabet is written in lower case.

Table 1 : Symbols of some elements based on their Latin names

Element	Symbol	Latin name
Gold	Au	Aurum
Silver	Ag	Argentum
Mercury	Hg	Hydrargyrum
Copper	Cu	Cuprum
Lead	Pb	Plumbum
Iron	Fe	Ferrum
Sodium	Na	Natrium
Potassium	K	Kalium
Tin	Sn	Stannum
Antimony	Sb	Stibium
Tungsten	W	Wolfram

1.2. CHEMICAL SYMBOL

A chemical symbol is an abbreviation or a short representation used to name the elements. Table 1 and 2 shows few examples of chemical symbols. It may be noted that in some elements first letter is same for more than one element. The symbols of such elements are formed from the first letter and another letter appearing later in the name. *e.g.* Carbon (C), Cobalt (Co), Chlorine (Cl), Copper (Cu) have first alphabet 'C' and thus symbols of Cobalt, Copper and Chlorine contains another letter along with 'C'.

Significance of Symbol

Symbol represents the following aspects :

(i) Name of the element, *e.g.* the symbol 'C' represents Carbon

(ii) One atom of element, *e.g.* the symbol 'C' represents one atom of carbon

(iii) Quantity of element (in mass) that is equal to its atomic mass, *e.g.* the symbol 'C' represents 12 parts by weight of Carbon, which is the atomic weight of carbon.

Table 2 : Symbols of some common elements

Hydrogen	H	Calcium	Ca
Lithium	Li	Potassium	K
Helium	He	Chromium	Cr
Boron	B	Nickel	Ni
Carbon	C	Zinc	Zn
Nitrogen	N	Cobalt	Co

Phosphorus	P	Copper	Cu
Fluorine	F	Iodine	I
Oxygen	O	Bromine	Br
Sodium	Na	Silver	Ag
Magnesium	Mg	Cadmium	Cd
Aluminium	Al	Arsenic	As
Silicon	Si	Neon	Ne
Sulphur	S	Barium	Ba
Chlorine	Cl	Beryllium	Be
Argon	Ar	Platinum	Pt

1.3. CHEMICAL FORMULA

Chemical formula of a compound is a symbolic representation of a particular composition of chemical compounds. It is represented by symbols of atoms and numbers. Molecules of compound contain atoms bonded in fixed ratio. *e.g.* K_2CO_3 represents one molecule of Potassium carbonate which contains 2 atoms of Potassium, 1 atom of Carbon and 3 atoms of Oxygen. Similarly, H_2SO_4 represents Sulphuric acid which contains 2 atoms of Hydrogen, 1 atom of sulphur and 4 atoms of Oxygen. Formula represents the ratio in which different atoms are present in that molecule of a compound.

Significance of formula

Formula represents the following aspects :

(i) Name of the compound, *e.g.* KNO_3 represents Nitric acid.

(ii) Number of various atoms present, *e.g.* Nitric acid is composed of 1 atom of Potassium, 1 atom of Nitrogen and 3 atoms of Oxygen.

(iii) Molecular mass of the compound, *e.g.* molecular mass of nitric acid is 101, given that atomic mass of Potassium is 39, Nitrogen is 14 and that of Oxygen is 16.

1.4. VALENCY

Valency of an element is defined as combining capacity of an atom. The number of electron that an atom can lose, gain or share during a chemical reaction is called its valency. Atoms lose or gain electrons to attain stability and form ions. An ion is a charged particle and can be positively or negatively charged.

An atom with a positively charged ion is called as cation. Cations are formed by losing electrons to attain stability. Elements with one, two or three electrons in the valence shell are usually metals.

For example—

$$\underset{(2,8,1)}{Na} - 1e^- \rightarrow \underset{(2,8)}{Na^+} \text{ (univalent)}$$

$$\underset{(2,8,1)}{Mg} - 2e^- \rightarrow \underset{(2,8)}{Mg^{2+}} \text{ (divalent)}$$

$$\underset{(2,8,1)}{Al} - 3e^- \rightarrow \underset{(2,8)}{Al^{3+}} \text{ (trivalent)}$$

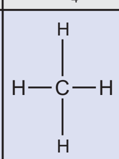
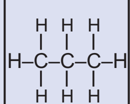
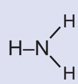
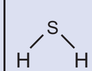
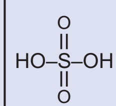
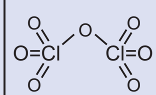
An atom with a negatively charged ion is called as anion. Anions are formed by gaining electrons to attain stability. Elements with five, six or seven electrons in the valence shell are known as non-metals.

$$\underset{(2,8,7)}{Cl} + 1e^- \rightarrow \underset{(2,8,8)}{Cl^-} \text{ (univalent)}$$

$$\underset{(2,6)}{O} + 2e^- \rightarrow \underset{(2,8)}{O^{2-}} \text{ (divalent)}$$

$$\underset{(2,5)}{N} + 3e^- \rightarrow \underset{(2,8)}{N^{3-}} \text{ (trivalent)}$$

Ions consist of one atom or group of atoms that carry a net charge on them. A group of atoms carrying a charge is called polyatomic ion.

Compound	H_2	CH_4	C_3H_8	C_2H_2	NH_3	NaCN	H_2S	H_2SO_4	Cl_2O_7
Diagram	H – H			H–C≡C–H		Na–C≡N			
Valencies	Hydrogen 1	Carbon, 4 Hydrogen, 1	Carbon, 4 Hydrogen, 1	Carbon, 4 Hydrogen, 1	Nitrogen, 3 Hydrogen, 1	Sodium, 1 Carbon, 4 Nitrogen, 3	Sulphur, 2 Hydrogen, 1	Sulphur, 6 Oxygen, 2 Hydrogen, 1	Chlorine, 7 Oxygen, 2

Variable Valency

Some elements combine with other atoms by donating, accepting or sharing electrons in different proportions depending on the nature of reaction. Some metals like Nickel, Copper, Tin and non-metals like Nitrogen, Phosphorus and Oxygen show variable valency. For elements exhibiting positive valencies, the suffix 'ous' is used for lower valency and the suffix 'ic' is used for higher valency. *e.g.* Cu_2O is Cuprous oxide, which is also written as Copper(I) oxide and CuO is Cupric oxide, which is also written as Copper(II) oxide. Roman number I and II in above examples indicates the valency of copper.

Table 3 : Examples of Variable Valency

Metal	Valency	Suffix used
Iron	2	Ferrous
	3	Ferric
Copper	1	Cuprous
	2	Cupric
Mercury	1	Mercurous
	2	Mercuric
Lead	2	Plumbous
	4	Plumbic
Tin	2	Stannous
	4	Stannic
Gold	1	Aurous
	3	Auric

1.5. RADICAL

A radical is an atom or a group of atoms of same or different elements that behaves as a single unit with a positive or negative charge. Based on the compound, radical can be classified as simple radical or compound radical.

Simple radicals are made up of one atom like Na^+, Al^{3+}, O^{2-}, etc.

Compound radicals are made up of group of atoms like Carbonate (CO_3^{2-}), Nitrate (NO_3^{1-}), Sulphate (SO_4^{2-}), etc.

Consider an example of neutralisation reaction, where an acid reacts with base to form salt and water.

$$NaOH + HCl \rightarrow NaCl + H_2O$$

In this reaction, Na^+ radical from the base and Cl^- radical from the acid is contributed to form sodium chloride. Therefore, Na^+ is a basic radical and Cl^- is an acidic radical. Each radical has its own combining capacity and we consider this while writing chemical formula.

Table 4 : List of some common electropositive ions (basic radicals)

Monovalent electropositive ions	Divalent electropositive ions	Trivalent and Tetravalent electropositive ions
Ammonium NH_4^{1+}	Barium Ba^{2+}	Aluminium Al^{3+}
Sodium Na^{1+}	Magnesium Mg^{2+}	Arsenic As^{3+}
Potassium K^{1+}	Calcium Ca^{2+}	Gold(III) Au^{3+}
Lithium Li^{1+}	Copper(II) Cu^{2+}	Bismuth Bi^{3+}
Hydrogen H^{1+}	Manganese Mn^{2+}	Chromium Cr^{3+}
Mercury (I) Hg^{1+}	Nickel Ni^{2+}	Ferric Fe^{3+}
	Zinc Zn^{2+}	Lead(IV) Pb^{4+}
	Lead(II) Pb^{2+}	Platinum(IV) Pt^{4+}
	Tin(II) Sn^{2+}	Tin(IV) Sn^{4+}
	Platinum(II) Pt^{2+}	

Table 5 : List of some common electronegative ions (acidic radicals)

Monovalent electronegative ions	Divalent electronegative ions	Trivalent and Tetravalent electronegative ions
Chloride Cl^{1-}	Carbonate CO_3^{2-}	Nitride N^{3-}
Acetate CH_3COO^{1-}	Oxide O^{2-}	Phosphide P^{3-}
Bicarbonate or hydrogen carbonate HCO_3^{1-}	Peroxide O_2^{2-}	Phosphate PO_4^{3-}
Bisulphide or hydrogen sulphide HS^{1-}	Sulphate SO_4^{2-}	Carbide C^{4-}
Bisulphate or hydrogen sulphate HSO_4^{1-}	Sulphite SO_3^{2-}	
Bisulphite or hydrogen sulphite HSO_3^{1-}	Sulphide S^{2-}	
Bromide Br^{1-}	Chromate CrO_4^{2-}	
Fluoride F^{1-}		
Iodide I^{1-}		
Cyanide CN^{1-}		
Nitrate NO_3^{1-}		
Nitrite NO_2^{1-}		
Permanganate MnO_4^{1-}		

INTEXT QUESTIONS

1. Explain each of the following terms with examples
   (a) Element (b) Atomicity
   (c) Valency (d) Chemical symbol
   (e) Chemical formula
2. Why do few elements show variable valency?
3. Find acidic and basic radicals in the following compounds
   (a) $NaClO_3$ (b) $MgSO_4$
   (c) $(NH_4)_2SO_4$ (d) $NaHCO_3$
   (e) $Mg(OH)_2$ (f) NaF
   (g) $KMnO_4$ (h) $Zn(NO_3)_2$
4. Write the chemical formulae of following compounds and then identify the acidic and basic radicals :
   (a) Calcium carbonate (b) Barium sulphate (c) Zinc chloride (d) Ammonium phosphate (e) Calcium phosphate (f) Calcium carbide (g) Mercury(II) oxide (h) Hydrogen peroxide (i) Iron(III) oxide (j) Chromium sulphate (k) Nickel bisulphate (l) Sodium cyanide (m) Aluminium oxide.

2. Few elements are said to exhibit variable valencies. Predict the valencies of metal ions in the given compounds and write their names
   (a) Fe_2O_3 (b) $CuCl_2$
   (c) HgO (d) $PbCl_4$
   (e) $SnCl_4$

1.6. WRITING CHEMICAL FORMULA

Chemical formula is a symbolic representation of chemical composition of a compound in terms of chemical symbols of elements, numbers and sometimes other symbols. The chemical formula of different compounds can be written using the knowledge of symbols and valencies of radicals.

Criss-cross method is one of the method used for this purpose. Following rules have to be followed while writing chemical formula–

(i) Symbols have to be written side by side. When a compound consist of a metal and a non-metal, the symbol of metal is written first. For example : In sodium chloride sodium is metal and chlorine is non-metal therefore, symbol of sodium is written first and then chlorine.Thus, the chemical formula of sodium chloride is NaCl, another example is magnesium chloride $MgCl_2$.

(ii) Valencies of each ion is written on the top of its symbol (superscript). Then the crossover of valencies of the combining atoms takes place. This number is written in the form of subscript. Positive and negative signs are ignored while crossing over the valencies.

(iii) If a compound contains polyatomic ions, the ions are enclosed in a bracket before writing the subscript number. However bracket is not required if the number of polyatomic ion is one.

(iv) Valency is divided by their Highest Common Factor (HCF), if any, to get a simple ratio.

Table 6 : Steps for writing chemical formulae

Name of the compound	Symbols with valencies	Criss-cross method	Formula
Magnesium chloride	Mg^{2+}, Cl^{1-}	Mg^{2+} Cl^{1-} Mg_1 Cl_2	$MgCl_2$
Aluminium chloride	Al^{3+}, Cl^{1-}	Al^{3+} Cl^{1-} Al_1 Cl_3	$AlCl_3$
Magnesium oxide	Mg^{2+}, O^{2-}	Mg^{2+} O^{2-} Mg_2 O_2 Cancel the common factor	MgO
Zinc hydroxide	Zn^{2+}, OH^{1-}	Zn^{2+} OH^{1-} Zn_1 OH_2	$Zn(OH)_2$
Zinc sulphate	Zn^{2+}, SO_4^{2-}	Zn^{2+} SO_4^{2-} Zn_2 $(SO_4)_2$ Cancel the common factor	$ZnSO_4$
Ammonium sulphate	NH_4^{1+}, SO_4^{2-}	NH_4^{1+} SO_4^{2-} $(NH_4)_2$ SO_4	$(NH_4)_2SO_4$
Iron(III) chloride	Fe^{3+}, Cl^{1-}	Fe^{3+} Cl^{1-} Fe_1 Cl_3	$FeCl_3$
Calcium nitrate	Ca^{2+}, NO_3^{1-}	Ca^{2+} NO_3^{1-} Ca_1 $(NO_3)_2$	$Ca(NO_3)_2$
Sodium carbonate	Na^{1+}, CO_3^{2-}	Na^{1+} CO_3^{2-} Na_2 CO_3	Na_2CO_3

Rules for Naming Certain Compounds

(i) Combination of metal and non-metal : The name of metal is written first followed by the non-metal with the suffix *'ide'*. *e.g.* Sodium Chloride.

(ii) Combination of two non-metals : The prefix *'tri'*, *'tetra'*, *'penta'*, etc. is written as per the requirement. *e.g.* Phosphorus trichloride PCl_5.

(iii) Naming acids : For binary acids, the prefix *hydro* and the suffix *ic* is used. *e.g.* Hydrochloric acid (HCl) or Hydrofluoric acid (HF). In case of acids containing polyatomic radicals, prefix is given based on the second element. *e.g.* Sulphuric acid ($H_2SO_4^{2-}$) where sulphur is the second element. For Nitric acid (HNO_3) nitrogen is the second element, while in Phosphoric acid (H_3PO_4) Phosphorus is the second element. Above given acids have their names based on their second element.

Some compounds do not follow any rule. Such names are called as common names or trivial names. e.g. dihydrogen oxide is called as water, while nitrogen trihydride is called as ammonia, etc.

1.7. CHEMICAL REACTION AND EQUATION

A chemical equation is a shorthand form to represent chemical change. A chemical equation shows the result of a chemical change, in which the ***Reactants*** and the ***Products*** are represented by symbols (in the case of element) and Formula (in the case of compounds).

(i) Reactant : The substance which take part in a chemical reaction.

(ii) Product : The substance formed as a result of chemical change.

Consider the following scenario of day to day life–

(i) Curdling of milk
(ii) Rusting of iron nail or window panes
(iii) Combustion of fuel
(iv) Digestion of food
(v) Fermentation of dough

In all the above situation the nature and properties of initial substance is different from that of the final product. Chemical reaction leads to a chemical change. Progress or completion of Chemical reaction can be characterised by any one of the following criteria :

(i) Change in state
(ii) Change in colour
(iii) Evolution of energy (in the form of light, heat)
(iv) Evolution of gas
(v) Formation of precipitate

When Iron nail reacts with copper sulphate, Iron(II) sulphate and copper are formed. This reaction can be written in the word equation form as shown below–

Iron + Copper sulphate → Iron(II) sulphate + Copper

Chemical reaction can be represented in a concise manner if we write chemical symbols and formulae instead of words. The above reaction can be written as :

$$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$$

A chemical equation can be defined as the symbolic representation of the chemical reaction using the symbols and formulae of the substances involved in the reaction.

A chemical reaction may involve :

(i) One reactant and more than two or more products.

(a) $CaCO_3 \xrightarrow{\Delta} CaO + CO_2$

(b) $KCl \longrightarrow 2K + Cl_2$

(ii) Two reactants and one product.

(a) $H_2 + O_2 \longrightarrow H_2O$

(b) $HCl + NH_3 \longrightarrow NH_4Cl$

(iii) Two reactants and two products.

(a) $NH_4OH + HCl \longrightarrow NH_4Cl + H_2O$

(b) $HCl + NaOH \longrightarrow NaCl + H_2$

(iv) Two reactants and three or more products

(a) $2\,KMnO_4 + 3H_2SO_4 \longrightarrow K_2SO_4 + 2MnSO_4 + 3H_2O + 5[O]$

(b) $CaCO_3 + HCl \longrightarrow CaCl_2 + H_2O + CO_2$

1.8. WRITING A CHEMICAL EQUATION

(i) Chemical formulae or symbols of reactants are written on the left hand side of the chemical equation. Each reactant is separated by a (+) sign.

(ii) Chemical formulae or symbols of products are written on the right hand side of the chemical equation. Each product is separated by a (+) sign.

(iii) The reactant and products are separated by an arrow sign (→) is written between the reactants and products.

(iv) Reactants and products are represented in their molecular forms and not atomic forms, as atomic forms are unstable.

Example—

$$\underbrace{Fe + CuSO_4}_{\text{Reactants}} \rightarrow \underbrace{FeSO_4 + Cu}_{\text{Products}}$$

In the above example, Iron reacts with Copper sulphate to give Iron(II) sulphate and Copper, where iron and copper sulphate are reactants whereas iron sulphate and copper are products.

1.9. BALANCING OF CHEMICAL EQUATIONS

As we know that the chemical equations are represented by the words, symbols and formulas. Here are some of the rules which needs to be followed while balancing a chemical equation.

1. In a balanced equation number of atoms of each element is same on the side of reactants and on the side of products.

2. All equations must be balanced to justify, "Law of Conversation of Mass", which states that the "matter is neither created nor destroyed in a chemical reaction."

3. An equation is generally balanced by **hit and trial method** or **Partial equation method.**

4. **Steps involved in balancing of chemical equation by Hit and Trial method**

(a) Determine the reactants and products in a chemical reaction.

(b) Count the number of atoms of each element on both sides of the equation.

(c) Select the elements that occur for the least number of times in the equation.

(d) Balancing of atoms of each element on both sides of the reaction.

(e) Always leave hydrogen and oxygen for last to balance.

(f) Balance the hydrogen atoms lastly followed by balancing of oxygen atoms.

Example :

Formation of Ammonia :

$$N_2 + H_2 \rightarrow NH_3 \quad \text{(Unbalanced equation)}$$

Step 1 : $N_2 + H_2 \rightarrow NH_3$

No. of atoms $N = 2 \;\; H = 2 \rightarrow N = 1, H = 3$

Step 2 : In the above equation number of Nitrogen atoms on both sides are not equal, multiply with suitable integer to balance the Nitrogen atoms on both sides. So, multiply with '2' on the product side.

$$N_2 + H_2 \rightarrow 2NH_3$$

In the above equation the number of Nitrogen atoms were balanced.

Step 3 : Hydrogen atoms on both sides were not balanced. So, multiply with suitable integer. Multiply with '3' to Hydrogen on reactant side.

Now, the above equation is balanced.

INTEXT QUESTIONS

Balance each of the following equations :

1. $NaNO_3 + PbO \rightarrow Pb(NO_3)_2 + Na_2O$
2. $Al + S_8 \rightarrow Al_2S_3$
3. $CH_4 + O_2 \rightarrow CO_2 + H_2O$
4. $NH_3 + O_2 \rightarrow N_2 + H_2O$
5. $NaOH + H_2CO_3 \rightarrow Na_2CO_3 + H_2O$

Importance of Chemical Equation

(i) Chemical equation tells us about the number of molecules of each substance participating in the chemical reaction.

(ii) Chemical equations can be made more informative by incorporating the details of reaction conditions. Information regarding the temperature at which reaction occurs, catalyst, pressure exerted, etc. can be written on or below the arrow.

(iii) The physical states of reactants and products can be written using letters (s) for solid, (l) for liquid, (g) for gas and (aq) for solution in water. For gas produced in the reaction it is '↑', for precipitate formed in the reaction it is '↓' and the direction of reaction is indicated by '→'.

For example—

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

However, a chemical equation cannot convey the following information about the chemical reaction :

(i) The time taken for the completion of the reaction.

(ii) Concentration of reactants and products.

(iii) The rate of reaction.

(iv) Whether the reaction is complete, incomplete or ongoing.

1.10. RELATIVE ATOMIC MASS (ATOMIC WEIGHT)

Atoms are the smallest indivisible part of the matter. Atoms of all elements differ in mass because of the different number of protons, electrons and neutrons they contain. The actual mass of an atom is extremely small. For example, the mass of single sulphur atom is approximately 5.3×10^{-24} and that of a hydrogen atom is 1.672×10^{-24}. Such small masses are not convenient for practical uses and calculations. Therefore, we use a relative atomic mass scale. In this scale an atom of carbon-12 is given a relative atomic mass of 12.00 while other atoms are given a relative atomic mass compared to that of carbon. Initially hydrogen was chosen as a unit and masses of other atoms were compared with it. Carbon-12 was selected because its adoption least affects the values of atomic masses of elements on the old standard.

The relative atomic mass or atomic weight of an element is the number of times one atom of the element is heavier than 1/12 times of the mass of an atom of carbon-12.

Above statement can be put up as an expression—

$$\boxed{\frac{\text{Mass of 1 atom of element}}{\text{1/12th the mass of one carbon - 12 atom}}}$$

Atomic mass is defined in atomic mass units [a.m.u.]. Atomic mass unit is defined as $1/12^{th}$ of the mass of carbon-12 atom.

Table 7 : Atomic masses of some common elements

Element	Symbol	Atomic mass (u)
Hydrogen	H	1
Carbon	C	12
Nitrogen	N	14
Oxygen	O	16
Sodium	Na	23
Magnesium	Mg	24
Sulphur	S	32
Chlorine	Cl	35.5
Calcium	Ca	40
Aluminium	Al	27
Copper	Cu	63.5
Iron	Fe	56
Potassium	K	39

1.11. RELATIVE MOLECULAR MASS (MOLECULAR WEIGHT)

The relative molecular mass of an element or a compound is the number that represents how many times one molecule of the substance is heavier than $1/12^{th}$ of the mass of an atom of carbon-12. This is obtained by adding together the relative atomic masses of all atoms present in that compound.

Observe few examples given below :

(a) Copper carbonate

Atomic masses (Cu-63.5, C-12, and O-16)

Chemical formula : $CuCO_3$

Mass of 1 copper atom = 63.5 amu

Mass of 1 carbon atom = 12 amu

Mass of 3 oxygen atom = 3 × 16= 48 amu

$= 63.5 + 12 + (3 \times 16)$

$= 123.5$

(b) Copper sulphate crystals

Atomic masses (Cu-63.5, S-32, O-16 and H-1)

Chemical formula : $CuSO_4.5H_2O$

$= 63.5 + 32 + (4 \times 16) + 5[(2 \times 1) + 16]$

$= 159.5 + 90$

$= 249.5$

(c) Calcium hydroxide

Atomic masses (Ca-40, O-16 and H-1)

Chemical formula : $Ca(OH)_2$

$= 40 + 2[16 + 1]$

$= 40 + 34$

$= 74$

(d) Acetic acid or ethanoic acid

Atomic masses (C-12, O-16 and H-1)

Chemical formula : CH_3COOH

$= (2 \times 12) + (4 \times 1) + (2 \times 16)$

$= 60$

1.12. PERCENTAGE COMPOSITION

Percentage composition of a compound is defined as the percentage by mass of each element present in it. It can be expressed as follows—

$$\boxed{\frac{\text{Total mass of the element in one molecule}}{\text{gram molecular weight of the compound}} \times 100}$$

Observe few examples given below—

(a) Calculate percentage composition of hydroen in sulphuric acid.

Molecular mass of H_2SO_4 is 98

Sulphuric acid contains 2 atoms of hydrogen, therefore, weight of 2 atoms of hydrogen is 2 g. Now, substitute these values in the formula.

$$\boxed{\frac{\text{Total weight of the element in one molecule}}{\text{gram molecular weight of the compound}} \times 100}$$

$$\frac{2}{98} \times 100 = 2.04 \text{ g of hydrogen}$$

$\therefore$ Hydrogen in sulphuric acid is 2.04%

(b) Calculate percentage composition of sodium in sodium bicarbonate

Molecular mass of $NaHCO_3 = 84$

$$\boxed{\frac{\text{Total weight of the element in one molecule}}{\text{gram molecular weight of the compound}} \times 100}$$

$$\frac{23}{84} \times 100 = 27.38 \text{ g of Sodium}$$

$\therefore$ Sodium in Sodium bicarbonate is 27.38%

(c) Calculate percentage composition of Hydrogen in copper sulphate crystals.

Molecular mass of $CuSO_4.5H_2O = 249.5$

$$\boxed{\frac{\text{Total weight of the element in one molecule}}{\text{gram molecular weight of the compound}} \times 100}$$

There are 5 molecules of water which contains 10 atoms of hydrogen.

$$\frac{10}{249 \cdot 5} \times 100 = 4.008 \text{ g of Hydrogen}$$

$\therefore$ Hydrogen in copper sulphate crystal is 4.008%.

We have seen through above examples that percentage composition determines composition abundance of the element in a molecule.

1.13. EMPIRICAL FORMULA OF A COMPOUND

The **empirical formula** of a compound is the simplest whole number ratio of each type of atom in a compound. It can be the same as the compound's molecular formula, but not always. An empirical formula can be calculated from information about the mass of each element in a compound or from the percentage composition.

Observe few examples given below :

Calculate empirical formula mass of following compounds :

(a) Hydrogen peroxide

Chemical formula : H_2O_2

Now, write the empirical formula for hydrogen peroxide in the simplest ratio.

Empirical formula will be HO as it is been divided by the highest common factor 2.

Now calculate empirical formula mass, HO = (1 + 16) = 17

(b) Butane

Chemical formula : C_4H_{10}

Now, write the empirical formula for butane in the simplest ratio.

Empirical formula will be C_2H_5 as it is been divided by the highest common factor 2.

Now, calculate empirical formula mass, C_2H_5 = $[(2 \times 12) + (5 \times 1)] = 29$

INTEXT QUESTIONS

1. Calculate the percentage of oxygen in water. (Given that the relative atomic mass of H= 1, O = 16)
2. Calculate the mass percentage of hydrogen present in water.
3. Find the percentage mass of water in washing soda crystals $Na_2CO_3.10H_2O$.
4. Calculate the molecular masses of the following molecules :
   (a) Nitric acid (HNO_3)
   (b) Glucose ($C_6H_{12}O_6$)
   (c) Water (H_2O)
   (d) Ammonia (NH_3)
5. Give empirical formula mass for the following :
   (a) Water
   (b) Sulphuric acid
   (c) Urea
   (d) Ammonia

SUMMARY

- **Chemistry** is the branch of science which deals with study of matter, investigation of their properties and the types of reactions they undergo to form new products.
- **Matter** is defined as anything that has mass and occupies space, which are made up of atoms, the smallest particles.
- An **element** is a simple and pure form of matter which cannot be decomposed into simpler substances.
- The number of atoms bonded in a molecule of an element is called its **atomicity**. Molecules of elements can be classified by its atomicity as Monoatomic molecule, diatomic molecule, Triatomic molecule, Tetratomic molecule and polyatomic molecule
- A **chemical symbol** is an abbreviation or a short representation used to denote the elements.
- **Chemical formula** of a compound is a symbolic representation of its composition.
- **Valency** is defined as combining capacity of an atom. The number of electron that an atom can lose, gain or share during a chemical reaction is called its valency.
- **Cations** are formed by losing electrons to attain stability. **Anions** are formed by gaining electrons to attain stability.
- A **chemical equation** can be defined as the symbolic representation of the chemical reaction using the symbols and formulae of the substances involved in the reaction. Since matter is neither created nor destroyed during a chemical reaction, every chemical equation must be balanced.
- The **relative atomic mass** or **atomic weight** of an element is the number of times one atom of the element is heavier than $1/12^{th}$ times of the mass of an atom of carbon-12.
- The **relative molecular mass** of an element or a compound is the number that represents how many times one molecule of the substance is heavier than $1/12^{th}$ of the mass of an atom of carbon-12. This is obtained by adding together the relative atomic masses of all atoms present in that compound.
- **Percentage composition** of a compound is defined as the percentage by weight of each element present in it. Percentage composition determines composition abundance of the element in a molecule.
- The **empirical formula** of a compound is the simplest formula, which gives the simplest positive integer ratio in whole numbers of atoms of different elements present in one molecule of the compound.
- **Empirical formula mass** is the sum of atomic masses of various elements present in the empirical formula.

EXERCISE

A. VERY SHORT ANSWER TYPE QUESTIONS

1. What is a chemical equation ?
2. Name two elements with their symbols, derived from their Latin names.
3. Define valency.
4. What are Radicals ?
5. Name two elements whose symbols are derived from its German name.
6. State the fundamental law that is applied in every chemical equation.

B. SHORT ANSWER TYPE QUESTIONS

1. Identify the acidic and basic radicals in below given example of where hydrochloric acid reacts with sodium hydroxide to form salt and water. Justify your answer
   $NaOH + HCl \rightarrow NaCl + H_2O$
2. Explain the criss-cross method of writing chemical formula with an example.
3. How are univalent, divalent and trivalent ions formed ? Give appropriate examples to support your answer.
4. State any three significances of chemical equations.
5. State any three drawbacks of chemical equations.
6. List out any three chemical reactions that can be observed in day today life.
7. List out the criteria that indicates the completion of the chemical reaction.
8. How are physical states of chemical reaction represented while writing a chemical equation ? Give an appropriate example to support your answer.

C. LONG ANSWER TYPE QUESTIONS

1. Why do some elements have variable valency give examples in support of your answer ?
2. (i) Define the terms with their formula :
   (a) Relative Atomic masses
   (b) Relative Molecular masses
   (ii) Define mass percentage, percentage composition and empirical formula.
3. What is the formula of urea ? Calculate the percentage of each element in its formula.
4. Why should an equation be balanced ? Explain with the help of a simple equation, give the importance of a chemical equation.

D. FILL IN THE BLANKS

(i) 1. Electronic configuration of Aluminium is ______.
2. The number of atoms bonded in a molecule of an element is called its _________.
3. Variable valency shown by Tin are _____ and _______.
4. ________ is a charged particle and can be positively or negatively charged.
5. Dalton use symbol _____ for hydrogen.

[**Ans :** **(i)** **1.** 2, 8, 3, **2.** valency, **3.** 2, 4, 4. Ion, **5.**]

(ii) Complete the following table

Anions → Cations ↓	Sulphate	Nitrate	Carbonate	Phosphate	Chloride	Oxide
Calcium	$CaSO_4$	$Ca(NO_3)_2$	$CaCO_3$	$Ca_3(PO_4)_2$	$CaCl_2$	CaO
Zinc						
Iron(III)						
Aluminium						
Sodium						
Potassium						
Ammonium						
Mercury(II)						
Copper(II)						

E. TRUE AND FALSE

1. Symbol for Oxygen is Oxy.
2. Atomic mass of carbon atom is 12 u.
3. Balancing of chemical equation is not essential to write an equation.
4. Formula for methane molecule is CH_4.
5. Chemical equation gives an idea about the time taken for the reaction to get completed.
6. Symbol for Calcium is Ca.

[Ans. 1. (F), 2. (T), 3. (F), 4. (T), 5. (F), 6. (T)]

F. MATCH THE COLUMN

Column A	Column B
1. Manganese	(i) Pb
2. Tin	(ii) O
3. Lead	(iii) Mn
4. Oxygen	(iv) Sn
5. Sulphur	(v) K
6. Sodium	(vi) Na
7. Potassium	(vii) S

[Ans. 1.→(iii), 2.→(iv), 3.→(i), 4.→(ii), 5.→(vii), 6.→(vi), 7.→(v)]

G. MULTIPLE CHOICE QUESTIONS

1. Formula of Ammonia is :
   (a) NH_3 (b) NH_2
   (c) CO_2 (d) N_3H
2. Modern atomic symbols are based on the method proposed by :
   (a) Bohr (b) Thomson
   (c) Dalton (d) Berzelius
3. CO_3^{2-} is the formula of which radical :
   (a) Carbonate (b) Chloride
   (c) Calcium (d) Oxide
4. Empirical formula of Benzene is :
   (a) C_6H_6 (b) CH_2
   (c) $C_6H_{12}O_6$ (d) NH_3
5. Variable valency shown by tin are :
   (a) 1, 2 (b) 1, 3
   (c) 2, 4 (d) 6, 8
6. Symbol for Cobalt is :
   (a) CO (b) Co
   (c) CO_1 (d) Cob

[Ans. 1. (a), 2. (d), 3. (a), 4. (a), 5. (c), 6. (b)]

H. OTHER QUESTIONS

Translate the following statements into chemical equations and then balance them :

1. Hydrogen gas reacts with nitrogen gas to form ammonia.
2. Sodium chloride reacts with aluminium sulphate to give sodium sulphate and ammonium chloride.
3. Potassium reacts with water to give potassium hydroxide and hydrogen gas.
4. Sodium bicarbonate reacts with hydrochloric acid to give sodium chloride, water and Carbondioxide.
5. Chlorine reacts with potassium bromide yielding potassium chloride and bromine.
6. Hydrogen sulphide burns in the atmosphere of oxygen to give water and sulphur dioxide.

I. OTHER QUESTIONS

Write the balanced chemical equation for the following reactions :

1. Barium chloride + Magnesium sulphate → Barium sulphate + Magnesium chloride
2. Mercury(II) sulphide + Ammonium nitrate → Mercury(II) nitrate + Ammonium sulphide
3. Sodium phosphate + Water → Sodium hydroxide + Phosphoric acid
4. Zinc phosphate + Hydrogen chloride → Zinc chloride + Phosphoric acid
5. Copper(II) nitrate + Silver → Copper + Silver nitrate
6. Iron oxide + Carbon monoxide → Iron + Carbondioxide
7. Sodium Carbonate + Sulphuric acid → Sodium sulphate + Water + Carbondioxide
8. Sodium carbonate + Silver nitrate → Sodium nitrate + Silver carbonate

9. $AgNO_3 + CuCl_2 \rightarrow AgCl + Cu(NO_3)_2$

10. $H_2C_2O_4 + KOH \rightarrow K_2C_2O_4 + H_2O$

11. $C_2H_2Cl_4 + Ca(OH)_2 \rightarrow C_2HCl_3 + CaCl_2 + H_2O$

12. $(NH_4)_2Cr_2O_7 \rightarrow N_2 + Cr_2O_3 + H_2O$

13. $Zn_3Sb_2 + H_2O \rightarrow Zn(OH)_2 + SbH_3$

14. $HClO_4 + P_4O_{10} \rightarrow H_3PO_4 + Cl_2O_7$

15. $C_6H_5Cl + SiCl_4 + Na \rightarrow (C_6H_5)_4Si + NaCl$

16. $Sb_2S_3 + HCl \rightarrow H_3SbCl_6 + H_2S$

17. $IBr + NH_3 \rightarrow NI_3 + NH_4Br$

18. $KrF_2 + H_2O \rightarrow Kr + O_2 + HF$

19. $Na_2CO_3 + C + N_2 \rightarrow NaCN + CO$

20. $K_4Fe(CN)_6 + H_2SO_4 + H_2O \rightarrow K_2SO_4 + FeSO_4 + (NH_4)_2SO_4 + CO$

✦ ✦ ✦

CHAPTER 2

CHEMICAL CHANGES AND REACTIONS

LEARNING OUTCOMES

2.1. INTRODUCTION

In our daily life, we come across many of changes. Some of these changes are temporary while some are permanent changes. These changes are broadly classified into two types, physical changes and chemical changes.

Physical changes are temporary changes where no new substance is formed for example, melting of ice, evaporation of water, etc.

A chemical change is permanent change where a new substance is formed for example combustion, rusting of iron, digestion of food etc.

2.2. CHEMICAL REACTION

Chemical reaction is a process of formation of products from reactants. During a chemical reaction, chemical bonds between atoms are formed or broken. The substances that undergo a chemical reaction are called the **reactants**, and the substances produced at the end of the reaction are known as the **products**. A chemical reaction is written in the form of chemical equations. A **chemical equation** can be defined as the symbolic representation of the chemical reaction using the symbols and formulae of the substances involved in the reaction. The reactants are written on the left hand side whereas products formed are written on the right hand side connected by one headed or two headed arrows. For example, a reaction

$$A + B \rightarrow C + D$$

Here, A and B are the reactants, which reacts to form the products C and D.

(a) Characteristics of a Chemical Reaction :

As we observe the changes around us, we can see that there is a large variety of chemical reactions taking place around us. Any of the following observations helps us to determine whether a chemical reaction has taken place–

(i) Change in state : Some reactions are characterized by change in state of reactants.

Examples :

(a) Combustion of petrol forms a water vapour and carbon dioxide which are in gaseous state.

$$2C_8H_{18\,(l)} + 25O_{2\,(g)} \rightarrow 16\ CO_{2\,(g)} + 18\ H_2O_{(l)}$$

(b) Hydrogen gas reacts with oxygen gas to give water

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

(ii) Evolution of a gas : Some reactions proceed by evolving a gas as a result of chemical reaction.

Examples :

(a) Zinc reacts with dilute sulphuric acid to give zinc sulphate and evolves hydrogen gas with effervescence.

$$Zn + dil.\ H_2SO_4 \rightarrow ZnSO_4 + H_2\uparrow$$

(b) Sodium bicarbonate reacts with dil. hydrochloric acid to give sodium chloride, water and carbon dioxide gas is evolved with effervescence

$$NaHCO_3 + dil.\ HCl \rightarrow NaCl + H_2O + CO_2\uparrow$$

(iii) Change in colour : Certain chemical reaction depicts a characteristic change in colour as the reaction proceeds.

Examples :

(a) Blue colour copper sulphate solution turns green when iron reacts with it.

$$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$$

(b) Silver nitrate reacts with sodium bromide to form a yellow precipitate of silver bromide.

$$AgNO_3 + NaBr \rightarrow AgBr + NaNO_3$$

(iv) Formation of precipitate : Some chemical reactions takes place with the formation of a precipitate. Precipitate is an insoluble solid that emerges from a liquid solution.

Examples :

(a) Sodium sulphate reacts with barium chloride to give sodium chloride and a white precipitate of barium sulphate.

$$Na_2SO_4 + BaCl_2 \rightarrow NaCl + BaSO_4$$

(b) Silver nitrate reacts with potassium chloride to give potassium nitrate and a white precipitate of silver chloride.

$$AgNO_3 + KCl \rightarrow AgCl + KNO_3$$

2.3. TYPES OF CHEMICAL REACTIONS

Chemical reactions are mainly classified in the following groups :

(A) Direct Combination Reactions

When two or more reactants combine to form single product, or two or more elements or compounds react chemically to form a single compound such reactions are known as combination reaction. Combination reactions can be represented as :

$$X + Y \rightarrow XY$$

Activity 1 : Burning of Magnesium Ribbon

(a) Take a piece of magnesium ribbon and rub it with sand paper to remove extra dirt from its surface.

(b) Hold the magnesium ribbon with a pair of tongs, light the magnesium and collect the ash of burnt ribbon in a china dish.

(c) It burns brilliantly and a white ash is left behind, which is of magnesium oxide.

(d) It has been observed that, in this reaction, magnesium and oxygen combines to form a product magnesium oxide.

Chemical reaction :

$$2Mg + O_2 \rightarrow 2MgO$$

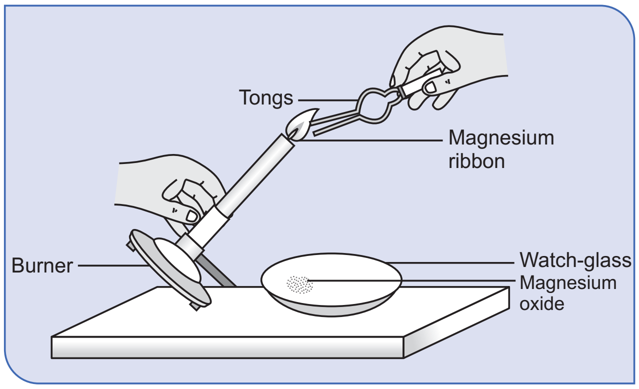


Fig. 2.1 : Burning of Magnesium Ribbon

Some other examples of combination reactions are :

(a) Reaction of hydrogen and chlorine to form hydrogen chloride. Here two reactants hydrogen and chlorine combines to form a product (hydrogen chloride).

$$\underset{\text{Hydrogen}}{H_2} + \underset{\text{Chlorine}}{Cl_2} \rightarrow \underset{\text{Hydrogen chloride}}{2HCl}$$

(b) Combination reaction of calcium and oxygen to form calcium oxide.

$$2Ca + O_2 \rightarrow 2CaO$$

(B) Decomposition Reaction

The reaction in which a compound is broken into two or more elements or smaller compounds are called as decomposition reaction. Decomposition reaction can be represented as follows :

$$XY \rightarrow X + Y$$

In these reactions, chemical bonds are broken, hence requires energy in the form of heat, light or electricity.

Some of the examples of decomposition reactions are :

(a) Zinc carbonate on heating gives zinc oxide and carbon dioxide gas is released.

$$ZnCO_3 \rightarrow ZnO + CO_2$$

(b) When limestone (calcium carbonate) is heated, it gives calcium oxide and carbon dioxide.

$$\underset{\substack{\text{Calcium}\\\text{carbonate}}}{CaCO_{3(s)}} \xrightarrow{\text{Heat}} \underset{\substack{\text{Calcium}\\\text{oxide}}}{CaO_{(s)}} + \underset{\substack{\text{Carbon}\\\text{dioxide}}}{CO_{2(g)}}$$

Decomposition reactions may be of three types :

(i) Thermal decomposition.

(ii) Electrolytic decomposition or electrolysis or electrochemical reaction.

(iii) Photo-decomposition or photochemical reaction

(i) Thermal decomposition : The chemical reaction where a single substance breaks into two or more simple substances by absorbing heat is called thermal decomposition. **For example :**

When heated strongly, potassium chlorate decomposes into potassium chloride and oxygen.

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

Activity 2 : Thermal Decomposition of Copper Carbonate

(a) Take 0.5 gram of copper carbonate in a test tube.

(b) Fit the test tube with a cork and delivery tube and put other end of the delivery tube into the test tube filled with lime water.

(c) Heat the test tube contents having copper carbonate over the flame, on heating the contents, a colourless gas carbon dioxide is given out with the formation of black coloured copper oxide.

(d) Allow this gas to pass through lime water, the lime water turns milky. This indicates that the gas evolved is carbon dioxide.

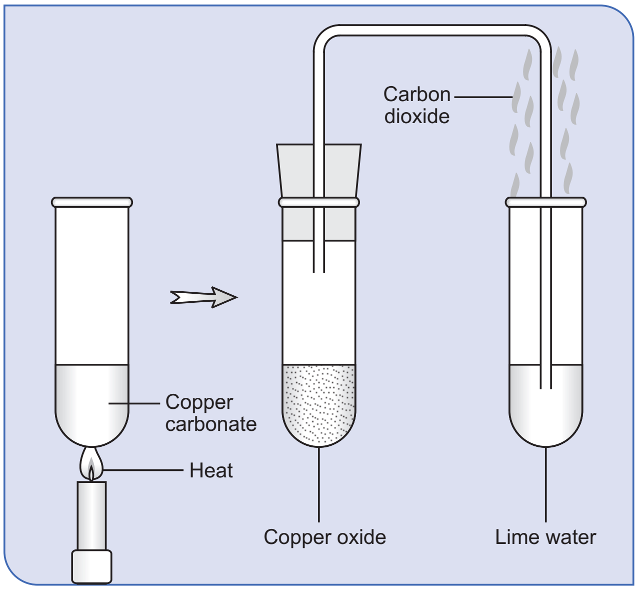


Fig. 2.2 : Thermal decomposition of copper carbonate

Thus, it is evident that copper carbonate decomposes to give carbon dioxide and copper oxide and can be represented as :

$$CuCO_3 \rightarrow CuO + CO_2$$

(ii) Electrolytic decomposition or electrolysis : The electrolytic decomposition reaction takes place by passing an electric current through the aqueous solution of a compound. Some common examples are :

Electrolysis of sodium chloride : On passing electric current through molten sodium chloride, it decomposes into sodium metal and chlorine gas.

$$2NaCl \rightarrow 2Na + Cl_2$$

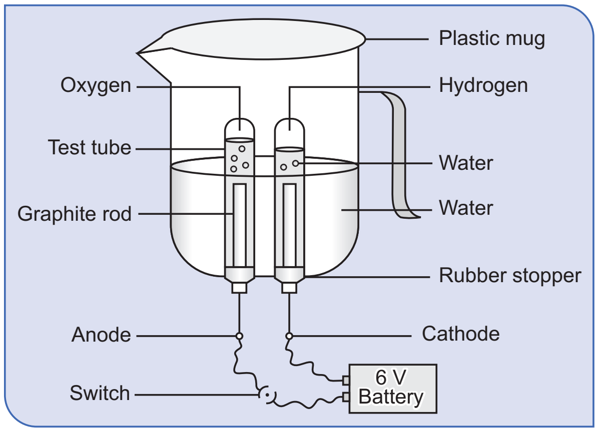


Fig. 2.3 : Electrolysis process

Electrolysis of aluminium oxide : On passing electricity through molten aluminium oxide, it decomposes into aluminium metal and oxygen gas.

$$2Al_2O_3 \rightarrow 4Al + 3O_2$$

Electrolysis of lead bromide : When an electric current is passed through molten lead bromide, it decomposes to lead metal and bromine gas.

$$PbBr_2 \rightarrow Pb + Br_2$$

Electrolysis of water : When an electric current is passed through acidulated distilled water, it decomposes to hydrogen and Oxygen.

$$2H_2O_{(l)} \xrightarrow{\textit{electricity}} 2H_{2(g)} + O_{2(g)}$$

(iii) Photo-decomposition or Photolysis : The decomposition reaction which occurs in the presence of sunlight is called photolysis or photo-decomposition reaction. These reactions take place with the absorption of light. Some of the examples of photo-decomposition reactions are :

(a) Decomposition of silver bromide : In the sunlight, silver bromide decomposes into silver and bromine gas.

$$\underset{\text{Light yellow}}{2AgBr_{(s)}} \xrightarrow{\text{sunlight}} \underset{\text{Grey}}{2Ag_{(s)}} + Br_{2(g)}$$

(b) Decomposition of hydrogen peroxide : In the presence of light, hydrogen peroxide decomposes into water and oxygen

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

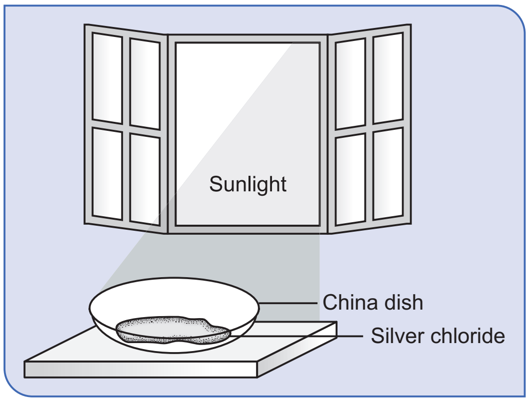


Fig. 2.4 : Decomposition of silver chloride

(c) Photosynthesis

$$6CO_2 + 12\,H_2O \xrightarrow{\text{sunlight}} C_6H_{12}O_6 + 6H_2O + 6O_2$$

Activity 3 : Decomposition of Silver Chloride

(a) Take 2 g of silver chloride in a China dish and note its white colour.

(b) Keep the China dish in direct sunlight.

(c) After some time, you will notice that the powder turns into grey. This happens because silver chloride decomposes into silver and chlorine.

$$2AgCl_{(s)} \xrightarrow{\text{sunlight}} 2Ag_{(s)} + Cl_{2(g)}$$

One of the major application of Photo-decomposition reaction is in the field of photography. Silver chloride is utilized in photographic paper. Photographic paper is coated with silver chloride. When a photographic paper is exposed to light, it turns grey due to photo-decomposition of silver chloride. This helps in getting an image on photographic paper.

INTEXT QUESTIONS

1. Why potassium iodide is kept in coloured bottles?
2. Which reaction is used in Photography?
3. Which reaction is used in laboratory preparation of oxygen?
4. Give reaction what happens when :
   (i) Ammonia reacts with hydrochloric acid.
   (ii) Iron reacts with sulphur.
   (iii) Potassium chlorate is heated
   (iv) Mercury(II) oxide is heated.
5. Which colour is obtained when white silver chloride is kept in Sunlight?
6. Explain decomposition reaction with example.
7. Hydrogen and chlorine reacts to form hydrogen chloride. Which type of reaction is this ?

(c) Displacement Reactions

When an atom or ion is removed or replaced by another atom or ion to form product, the reaction is called displacement reaction. Displacement reaction is also called substitution reaction. In displacement reaction, more reactive element displaces a less reactive element from its compound. These reactions generally take place in solutions and releases energy; thus they are exothermic reactions. The general equation for displacement reaction is written as follows :

$$XY + Z \rightarrow XZ + Y$$

Some of the examples of displacement reactions are :

(i) When Zinc granules react with copper sulphate solution, zinc displaces copper from copper sulphate solution to form zinc sulphate.

$$Zn_{(s)} + CuSO_{4(aq)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$$

(ii) When lead reacts with copper chloride solution, lead displaces copper from copper chloride solution to form lead chloride.

$$Pb_{(s)} + CuCl_{2(aq)} \rightarrow PbCl_{2(s)} + Cu_{(s)}$$

(iii) Iron displaces lead from lead nitrate solution in all the above reactions, more reactive metals

$$Fe_{(s)} + Pb(NO_3)_{2(aq)} \rightarrow Fe(NO_3)_{2(aq)} + Pb_{(s)}$$

displaces less reactive metal. The relative reactivity of metals or non-metals can be understood from

reactivity series. The metals are arranged in order of their decreasing reactivity, this series is called reactivity series or activity Series. The metal placed higher in the activity series can displace the metal that occupies a lower position from the aqueous solution of its salt. Thus, potassium is the most reactive metal and Gold is the least reactive metal. All the metals above hydrogen are more reactive, they can displace hydrogen from acids or water and they can react with acids to give out hydrogen, whereas metals below hydrogen are less reactive, they cannot replace hydrogen from acids and thus they do not react with acids to give out hydrogen.

Reactivity Series of Metals

	Metal	Symbol	
These metals are more reactive than hydrogen	Potassium	K	**(Most reactive metal)**
	Sodium	Na	
	Calcium	Ca	
	Magnesium	Mg	
	Aluminium	Al	
	Zinc	Zn	
	Iron	Fe	
	Tin	Sn	
	Lead	Pb	
	[Hydrogen]	[H]	
These metals are less reactive than hydrogen	Copper	Cu	
	Mercury	Hg	
	Silver	Ag	
	Gold	Au	**(Least reactive metal)**

(iv) Double decomposition reaction or Double displacement reaction : The reaction in which there is an exchange of ions between two reactants is called double displacement reaction. These may be defined as the chemical reactions in which one component each from both the reacting molecules is exchanged to form the products. During this reaction, the cations and anions of two different compounds are exchanged, and forms two entirely different compounds. It can be represented as :

$$PQ + RS \rightarrow PR + QS$$

Some of the examples of double decomposition reaction are :

(a) Iron sulphide reacts with dilute sulphuric acid.

$$FeS_{(s)} + H_2SO_{4(aq)} \rightarrow FeSO_{4(aq)} + H_2S_{(g)}$$

(b) Reaction of copper sulphate and ammonium hydroxide.

$$CuSO_4 + 2NH_4OH \rightarrow Cu(OH)_2 + (NH_4)_2SO_4$$

(c) Reaction of silver nitrate and sodium chloride,

$$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl + NaNO_{3(aq)}$$

Activity 4 : Reaction of Barium Chloride and Sodium Sulphate.

(a) Take about 3 mL of sodium sulphate solution in one test tube and about 3 mL of barium chloride solution in another test tube.

(b) Pour sodium sulphate solution in the test tube which contains barium chloride solution.

Observations : A white precipitate of barium sulphate appears in the test tube.

$$\underset{\text{Sodium sulphate}}{Na_2SO_{4(aq)}} + \underset{\text{Barium chloride}}{BaCl_{2(aq)}} \rightarrow \underset{\text{Barium sulphate}}{BaSO_{4(s)}} + \underset{\text{Sodium chloride}}{2NaCl_{(aq)}}$$

Since in the above reaction, anions and cations are exchanged. It is an example of double decomposition reaction.

Double displacement reactions can be further classified as :

(i) Neutralization reactions

(ii) Precipitation reactions

(iii) Gas formation reactions

(i) Neutralization Reactions : When an acid and a base reacts to form salt and water, the reaction is known as neutralization reaction. These reactions are a specific kind of double displacement reaction. It is an acid-base reaction.

$$\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water}$$

Some examples of neutralization reactions are :

(a) Reaction of hydrochloric acid and sodium hydroxide.

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

(b) Reaction of potassium hydroxide and hydrochloric acid.

$$KOH + HCl \rightarrow KCl + H_2O$$

(c) Reaction of ammonium hydroxide and hydrochloric acid.

$$NH_4OH + HCl \rightarrow NH_4Cl + H_2O$$

(ii) Precipitation reaction : Precipitation is the formation of a solid in a solution or inside another solid during a chemical reaction. Some of the examples of Precipitation reactions are :

(a) On mixing aqueous solutions of silver nitrate and sodium chloride, a white curdy precipitate of silver chloride is formed.

$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

(b) On mixing an aqueous solution of lead nitrate with sodium sulphate, a white precipitate of lead sulphate is formed.

$PbNO_{3(aq)} + Na_2SO_{4(aq)} \rightarrow PbNO_{4(s)} + 2NaNO_{3(aq)}$

(c) On adding aluminium sulphate solution to calcium chloride solution, a precipitate of calcium sulphate is formed.

$Al_2(SO_4)_{3(aq)} + 3Ca(OH)_{2(aq)} \rightarrow 2Al(OH)_{3(aq)} + 3CaSO_{4(s)}$

Activity 5 : To show the Precipitation Reaction

(a) Take a clean test tube A and pour about 5 ml of lead nitrate solution to it.

(b) Take another test tube B, and pour in it about 5 ml of Potassium iodide.

(c) Now mix the contents of both the test-tubes.

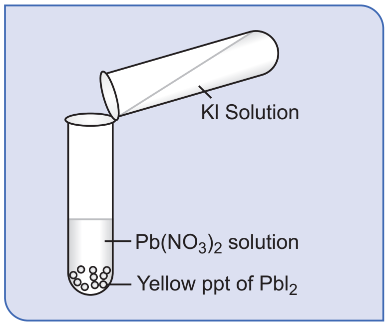


Fig. 2.5 : Precipitation reaction

Observation : A yellow precipitate is obtained. Thus, it is an example of Precipitation reaction.

$KI + Pb(NO_3)_2 \rightarrow KNO_3 + PbI_2$

(iii) Gas formation reaction : Sometimes a gas will be involved as one of the reactant or products in a solution reaction. For example :

$2HCl_{(aq)} + Na_2S_{(aq)} \rightarrow H_2S_{(g)} + 2NCl_{(aq)}$

INTEXT QUESTIONS

1. When dilute HCl is added to Copper oxide, the solution formed is bluish green. What is the product formed which is responsible for the colour. Give equation also.
2. Give equation what happens when,
   (i) Copper is added to zinc sulphate solution.
   (ii) Zinc is added to copper sulphate solution.
   (iii) Copper sulphate is added to hydrogen sulphide.
3. State whether the given reaction is feasible (possible) or not :
   (i) Reaction of silver and copper nitrate.
   (ii) Reaction of lead and copper chloride.
   (iii) Reaction of copper and hydrochloric acid.
4. Which reaction is used in Thermite Welding ?
5. Classify the reactions :
   (i) $CoCl_{2(aq)} + Na_2CO_{3(aq)} \rightarrow CoCO_{3\,(aq)} + 2NaCl_{(aq)}$
   (ii) $2AgCl_{(s)} \xrightarrow{\text{sunlight}} 2Ag_{(s)} + Cl_{2(g)}$
   (iii) $\underset{\text{Ferrous sulphate}}{2FeSO_{4(s)}} \xrightarrow{\text{Heat}} \underset{\text{Ferric oxide}}{Fe_2O_{3(s)}} + SO_{2(g)} + SO_{3(g)}$
   (iv) $\underset{\text{Magnesium}}{2Mg_{(s)}} + \underset{\text{Oxygen}}{O_{2(g)}} \longrightarrow \underset{\text{Magnesium oxide}}{2MgO_{(s)}}$
6. Arrange the following metals in increasing order of their reactivity. Zn, Pb, Au, Na, Ni.
7. When potassium hydroxide reacts with nitric acid, it gives potassium nitrate and water. Identify the reaction and give equation also.

2.4. ENERGY CHANGES IN A CHEMICAL REACTION

A chemical reaction always takes place either with absorption of energy or release of energy. In a chemical reaction, existing bonds are broken and new bonds are formed. Energy is absorbed when bonds are broken and released when bonds are formed. Thus, chemical reaction always involves change in energy. Every substance has a fixed amount of stored energy in the form of potential energy, which is referred to as chemical energy. In a chemical reaction energy is released or absorbed in the form of heat, light, or both. Based whether energy is absorbed or released, chemical reactions are of two types–

(i) Endothermic reactions

(ii) Exothermic reactions

(i) Endothermic reactions : A chemical reaction in which heat (in the form of energy) is absorbed is called endothermic reaction. Endothermic reaction leads to decrease in the temperature. Decomposition reactions are usually endothermic. Some of the examples of endothermic reactions are :

(a) Combination of nitrogen and oxygen to form Nitric oxide.

$$N_2 + O_2 + \text{Heat} \rightarrow 2NO$$

(b) The decomposition of water into hydrogen and oxygen.

$$2H_2O \rightarrow 2H_2 + O_2$$

(ii) Exothermic reaction : A chemical reaction in which heat (in the form of energy) is given out is called exothermic reaction. Excess energy from the reaction is released as heat, sound or light. Some other examples of exothermic reactions are :

(a) Sodium oxide dissolves in water to form sodium hydroxide and releases large amount of energy.

$$Na_2O + H_2O \rightarrow 2NaOH + \text{Energy}$$

(b) Carbon dioxide and water reacts to form carbonic acid.

$$CO_2 + H_2O \rightarrow H_2CO_3 + \text{Energy}$$

Difference between endothermic and exothermic reactions

S. No.	Endothermic reactions	Exothermic reactions
1.	Energy is absorbed in the endothermic reactions, from the surroundings.	Energy is released in exothermic reactions, in the surroundings.
2.	Energy of reactants is less than products.	Energy of reactants is more than products.
3.	Change in enthalpy is always positive.	Change in enthalpy is always negative.
4.	Free energy is positive.	Free energy is negative.
5.	Energy is released in the form of heat.	Energy is released in the form of heat, electricity, light or sound.
6.	**Examples :** Photosynthesis, cooking an egg etc.	**Examples :** Respiration, combustion etc.

INTEXT QUESTIONS

1. State whether the following reactions are exothermic or endothermic.
   (a) Respiration
   (b) Decomposition of limestone
   (c) Electrolysis of water
   (d) Burning of LPG gas.
2. What are exothermic reactions. Give examples.
3. Give some examples of reactions in which energy is absorbed.
4. Photosynthesis is an endothermic reaction. Explain.

SUMMARY

- Chemical reaction involves breaking and formation of bonds. A chemical reaction is written in the form of chemical equations.
- Chemical reactions are mainly classified in following groups :

  (a) Direct combination reaction : "When two or more reactants combine to form one product, or two or more elements or compounds react chemically to form a single compound is known as Combination reaction".

  (b) Decomposition reaction : These are the reaction in which one reactant gives two or more products after a reaction. These are opposite of combination reactions. Decomposition reactions may be of three types :

(i) **Thermal decomposition** : Decomposition takes place by heat.

(ii) **Electrolytic decomposition or Electrolysis** : Decomposition takes place by electricity.

(iii) **Photo-decomposition** : Decomposition takes place by light.

(c) Displacement Reaction : Displacement reaction is also called substitution reaction. It is a chemical reaction in which a more reactive element displaces a less reactive element from its compound. Displacement reactions can be classified as :

(a) Cation replacement reaction or metal replacement reactions.

(b) Anion replacement reaction or non-metal replacement reactions.

(d) Double Decomposition reaction or Double displacement reaction : The reaction in which there is an exchange of ions between two reactants is called double displacement reaction. Double displacement reactions can be further classified as

(i) Neutralization reactions

(ii) Precipitation reactions

(iii) Gas formation reactions

- A reaction which absorbs energy is called endothermic reaction.
- A reaction which releases energy is called exothermic reaction.

EXERCISE

A. VERY SHORT ANSWER TYPE QUESTIONS

1. Give an example of combination reaction in which two elements reacts ?

2. What is decomposition reaction ?

3. Give equation for electrolysis of water. What is ratio of oxygen and hydrogen liberated ?

4. Is respiration an exothermic reaction ? Explain.

5. What are displacement reactions ? Give example.

6. Silver chloride bottles are stored in coloured bottles. Why.

7. Why copper does not react with dilute sulphuric acid ?

8. What is the formula of quick lime and slaked lime ?

9. In the following chemical reaction, what is X ?

$$AlCl_3 + NaOH \rightarrow X + NaCl$$

10. What is a chemical reaction ?

11. Give an example of displacement reaction having nitrogen as one of the component.

12. Give an example of photochemical equation which involves hydrogen as component.

13. Give equation for electrolysis of water.

14. Photosynthesis is a photochemical reaction. Explain.

15. What happens if quick lime is added to water ?

16. Give an example of double displacement reaction.

17. What happens when potassium iodide is added to lead nitrate solution ?

18. Give an example where a metal replaces hydrogen from acids.

19. Which reaction takes place during digestion of food in our body ?

B. SHORT ANSWER TYPE QUESTIONS

1. Why energy released during respiration and burning is different ?

2. What do you mean by activity series of metals. Which element is at the top and which is at the bottom ?

3. Explain combination reactions with example.

4. Give an example of

(i) Combination of element with another element.

(ii) Combination of element with compound.

(iii) Combination of two compounds.

5. Classify the following reactions as endothermic and exothermic.

(i) Dilution of sulphuric acid

(ii) Decomposition of calcium carbonate

(iii) Respiration

6. When barium chloride is added to sodium sulphate, a precipitate is formed,

(a) What is the name and colour of precipitate ?

(b) Which type of reaction is this ?

(c) Write a balanced chemical equation for this.

7. AgCl is kept in sunlight for some time,
   (i) What colour change is observed in the reaction ?
   (ii) What is the reason for the colour change ?
   (iii) Give a balanced chemical equation and type of reaction.
8. What happens when copper wire is dipped in silver nitrate solution ?
9. Predict whether the following reactions are feasible or not ?
   (i) Silver plate is dipped in copper sulphate solution.
   (ii) Zn is added to dilute sulphuric acid.
   (iii) Copper is added to dilute hydrochloric acid.
10. Define neutralisation reaction with example.
11. In thermal decomposition of ferrous sulphate, precipitate is formed.
   (i) What is the colour of precipitate.
   (ii) Give a balanced chemical equation for the reaction.

C. LONG ANSWER TYPE QUESTIONS

1. Classify the reactions :
   (i) Digestion of food.
   (ii) Heating of limestone
   (iii) Lead reacts with copper chloride
   (iv) Magnesium is burnt in air
   (v) Respiration
2. Give an activity to illustrate thermal decomposition reaction.
3. How many type of decomposition reactions are there ? Give one example of each.
4. Differentiate between Exothermic and Endothermic reactions.
5. Compete the following reaction :
   (i) $NaOH + HCl \longrightarrow$
   (ii) $AgNO_3 + \longrightarrow AgCl +$
   (iii) $PbBr_2 \longrightarrow +$
   (iv) $CuSO_4 + Zn \longrightarrow Cu +$
   (v) $\quad + 3CO \longrightarrow 2Fe +$
6. What is the difference between Combination and decomposition reactions? Explain with examples.
7. Explain the types of double displacement reaction with examples.
8. Give equations :
   (a) Combination of ammonia and hydrochloric acid
   (b) Reaction of nitric oxide and oxygen
   (c) Iron and copper sulphate solution reacts with each other
   (d) Reaction of chlorine and potassium bromide solution.
   (e) Ammonium hydroxide and hydrochloric acid

D. FILL IN THE BLANKS

1. When quicklime reacts with water, it forms ....
2. .... reaction is when two or more simple molecules join to form a more complicated molecule.
3. Decomposition of calcium carbonate gives .... gas.
4. When decomposition takes place in the presence of light, it is called .... decomposition.
5. During electrolysis of water, ...gas is collected at anode.
6. Zinc is ....reactive than copper.
7. Precipitate of barium chloride is ....in colour.
8. Thermal decomposition is an example of the ....reaction.
9. The colour of copper sulphate solution is ....
10. Among the halogens, ...is most reactive.
11. ....metal is used in Thermite welding.
12. In a neutralisation reaction, acid and base reacts to form .... and water.

[**Ans** : 1. slaked Lime 2. combination 3. Carbon dioxide 4. Photolytic 5. oxygen 6. more 7. white 8. endothermic 9. blue 10. fluorine 11. Aluminium 12. salt.]

E. TRUE AND FALSE

1. The new substances formed as a result of a chemical reaction are called Reactants.
2. The formation of chemical bonds absorbs energy.

3. Quick lime is calcium oxide.
4. Electrolysis of aqueous sodium chloride is an example of thermal Decomposition reaction.
5. In photography, Photographic paper is coated with silver chloride
6. Iron is more reactive than copper.
7. Copper reacts with acids to liberate hydrogen gas.
8. Precipitation is a type of double decomposition reaction.
9. In the Reactivity Series, Potassium is least reactive and Gold is most reactive.
10. Digestion of food in our body is an example of decomposition reaction.

[**Ans** : 1. False 2. False 3. True 4. False 5. True 6. True 7. False 8. True 9. False 10. True]

F. MATCH THE COLUMN

	Column A		Column B
1.	Double Desplacement Reaction	(i)	$N_2 + 3H_2 \rightarrow 2NH_3$
2.	Displacement Reaction	(ii)	$ZnCO_3 \rightarrow ZnO + CO_2$
3.	Neutralization reaction	(iii)	$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
4.	Decomposition reaction	(iv)	$KOH + HCl \rightarrow KCl + H_2O$
5.	Combination reation	(v)	$Fe + ZnSO_4 \rightarrow ZnSO_4 + Fe$

[**Ans** : 1. (iii), 2. (v), 3. (iv), 4. (ii), 5. (i)]

G. MULTIPLE CHOICE QUESTIONS

1. Because of the formation of which of the following lime water turns milky when carbon dioxide is passed in it ?
(a) Calcium carbonate
(b) Calcium bicarbonate
(c) Calcium hydroxide
(d) Sodium carbonate

2. When copper slate crystals are dissolved in water, the colour of solution becomes :
(a) Green (b) Yellow
(c) Blue (d) Red

3. Which of the following reaction will not takes place ?
(a) Reaction of Al and Copper sulphate
(b) Reaction of Copper and copper sulphate
(c) Reaction of silver and hydrochloric acid
(d) Reaction of Iron and copper sulphate

4. The product obtained when sodium hydroxide reacts with hydrochloric acid ?
(a) Calcium chloride
(b) Sodium chloride
(c) Sodium hydride
(d) Hydrogen cyanide

5. Which of the following is an example of double displacement reaction ?
(a) Reaction of lead and copper chloride
(b) Reaction of Carbon and oxygen to form carbon dioxide
(c) Reaction of zinc and copper sulphate
(d) Reaction of sodium sulphate and barium chloride

6. Which salt is formed by neutralisation of sodium hydroxide and acetic acid ?
(a) Sodium acetate
(b) Sodium oxalate
(c) Sodium sulphate
(d) Sodium chloride

7. The chemical reaction between copper sulphate and iron nail is :
(a) Decomposition reaction
(b) Combination reaction
(c) Displacement reaction
(d) Double displacement reaction

8. Respiration is a :
(a) Endothermic reaction
(b) Exothermic reaction
(c) Both (a) and (b)
(d) None of these

9. A chemical reaction involves in :
(a) Only breaking of bonds
(b) Only formation of bonds
(c) Both breaking and formation of bonds
(d) None of these

10. When sulphuric acid reacts with zinc, which of the following gas is formed ?
(a) Sulphurdioxide
(b) Hydrogen
(c) Oxygen
(d) Zinc dioxide

11. What is the meaning of symbol "↓" represents in a chemical reaction ?
(a) Gas released
(b) Solid state
(c) Direction of reaction
(d) Precipitate formed

12. Which gas is formed by burning of natural gas ?
(a) Oxygen
(b) Carbon dioxide
(c) Hydrogen
(d) Ammonia

13. By decomposition of lead nitrate, yellow precipitate is formed. The precipitate is of
(a) Lead oxide
(b) Nitrogen dioxide
(c) Lead
(d) None of these

14. During electrolysis of molten sodium chloride, which gas is formed :
(a) Sodium
(b) Oxygen
(c) Nitrogen
(d) Chlorine

15. The correct order of reactivity is
(a) $F_2 > Cl_2 > Br_2 > I_2$
(b) $F_2 < Cl_2 < Br_2 < I_2$
(c) $F_2 < Cl_2 > Br_2 > I_2$
(d) $F_2 > Cl_2 < Br_2 < I_2$

16. What happens when silver plate is dipped in copper nitrate solution,
(a) Silver replaces copper
(b) Copper metal is deposited
(c) Colour of solution changes
(d) No reaction occurs

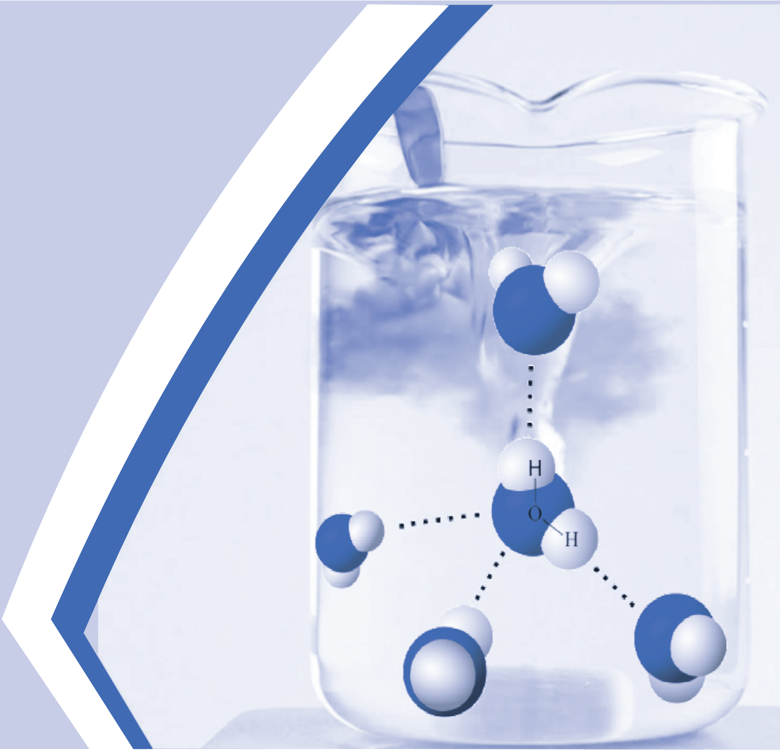
17. What is the general representation of displacement reaction ?
(a) $XY \longrightarrow X + Y$
(b) $X + Y \longrightarrow XY$
(c) $XY + Z \longrightarrow XZ + Y$
(d) $XY + ZA \longrightarrow XA + YZ$

[**Ans :** 1. (a) 2. (c) 3. (c) 4. (b) 5. (d) 6. (a) 7. (c) 8. (b) 9. (c) 10. (b) 11. (d) 12. (b) 13. (a) 14. (d) 15. (a) 16. (d) 17. (c)]

✦ ✦ ✦

CHAPTER 3

WATER



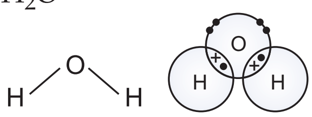
LEARNING OUTCOMES

3.1. INTRODUCTION

Water covers about 70% of Earth's surface, out of which 2.5% constitutes of fresh water. It is vital substance for all the living organisms. It is amazing to know that human body constitutes of about 70% water. Water occurs in both combined state as well as in free state. Typically, water exist in the liquid state. However, water boils at 100°C to form water vapours and freezes at 0°C to form ice. Water is found in combined state in all living beings. It is also present in hydrated salts in the form of water of crystallization. e.g. $CuSO_4.5H_2O$

Formula : H_2O

Structure



Biochemical Name : Dihydrogen oxide

Molecular Mass :

$$2 \times 1 + 16 \times 1 = 18 \text{ amu}$$

3.2. THE WATER CYCLE

The water cycle is the process by which water circulates between the, water bodies on the Earth's surface and the atmosphere and back to earth's surface as rain water.

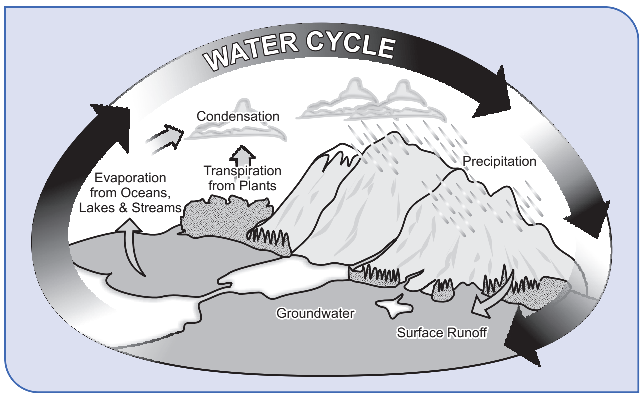


Fig. 3.1 : Water Cycle

3.3. PHYSICAL PROPERTIES OF WATER

1. Nature : Water is a colourless, odourless and a tasteless liquid.

2. Density : Density is defined as mass per unit volume. Therefore, the density of water is the weight of the water per its unit volume. Density depends on temperature, as the temperature increases density increases upto 4°C and then starts decreasing. Water has its maximum density 1 g/cm^3 or 1000 kg/m^3 at 4°C and density of 958.4 kilogram per cubic metre at 100°C. When water freezes, turning into ice, the density decreases. Due to lower density of ice (relative density 0.92) it floats.

3. Boiling point : Water boils and vaporises at 100°C at a pressure of 760 mm of Hg (normal pressure). However, this value is not constant. The boiling point of water depends on the atmospheric pressure, which changes according to elevation. Water boils at a lower temperature at the higher altitudes. The boiling point of water also depends on the purity of water. Water containing impurities (such as salted water) boils at a higher temperature than the pure water. When the water is boiled, it changes its state from a liquid to a gas. Water starts boiling and vaporizing at 100°Celsius. To change the phase of liquid water to water vapour it consumes heat. This heat is known as **latent heat of evaporation.**

4. Freezing point : Water freezes and solidifies at 0°C at a pressure of 760 mm of Hg (normal pressure). Freezing point decreases with the decrease in pressure. When water vapour condenses to water droplet it releases heat, such heat is known as **latent heat of condensation.**

5. Anomalous expansion of water : Generally, liquids expand on heating and contract on cooling. The anomalous expansion of water is an abnormal property of water where it expands when the temperature decreases from 4ºC to 0ºC and the density decreases. Thus, ice floats on the surface of the warm water. Thus, in colder regions marine life can survive under water when the top layer of water freezes.

6. Specific heat capacity : The capability of a molecule to absorb heat energy is called heat capacity. Water has capacity of absorbing 4.184 Joule heat at the temperature of 1 gram of water to increase 1 degree celsius (°C). Due to its high specific heat capacity water, can absorb a large amount of heat energy from the car engines. Therefore, water is used as coolant in motor vehicles.

7. Latent heat of vaporization of water : The amount of heat required to convert unit mass of a liquid into the vapour without a change in temperature is called **heat of vaporization**. Water in its liquid form has an unusually high boiling point close to 100°C. Because of the network of hydrogen bonding present in water molecules, a high input of energy is required to change 1 gram of liquid water to water vapour. The energy required for such a change, that is heat of vaporization of water is 40.65kJ/mol. A considerable amount of heat energy (586 calories) is required to accomplish such a change in water. This process occurs on the surface of water. Below the boiling point of water, the individual molecules of water acquire enough energy from each other in such a way that some surface water molecules can escape and vaporize. This process is known as evaporation.

8. Latent heat of fusion : The latent heat of fusion is the quantity of thermal energy required for a specific quantity of the substance to change its state from a solid to a liquid at constant pressure and vice versa. For example, 336 J/g is the specific latent heat of fusion of ice that is the amount of heat required to change 1 kg of ice to water without a change in temperature. Similarly, while ice melts, and the liquid water is formed with the latent heat of fusion remains at 0°C. Therefore, the heat of fusion for water at 0°C is approximately 334 joule (79.7 calorie) per gram.

3.4. SOLUBILITY OF WATER

Water is called the "universal solvent" because of its capacity to dissolve more substances than any other liquid. Water is a polar molecule and has high dielectric constant. Therefore, it reduces electrostatic force of attraction between positive and negative ions. Water can also dissolve other liquids and gases.

Experiment 1 : To show water contains dissolved solids

Procedure :

Set up a Bunsen burner, tripod stand and wire as shown in the figure. Place a beaker containing water over it. Take a watch-glass and put some tap water on it and keep it over the beaker. Heat the water in the beaker until it boils, and then let it boil briskly. When all the water on the watch-glass has evaporated, turn off the Burner and remove the watch-glass with a pair of tongs and let it cool. Examine the watch-glass for traces of solid residue.

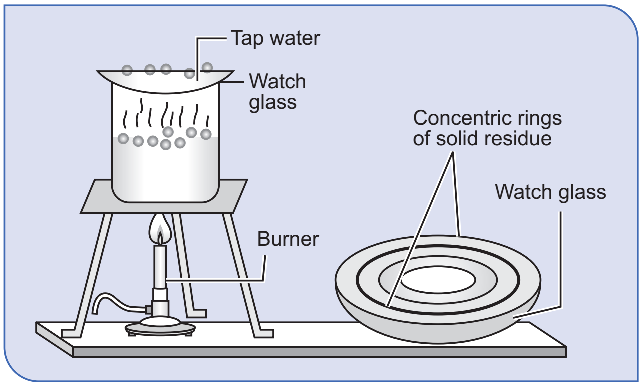


Fig. 3.2 : To show water contains dissolved solids

Observation :

When the tap water is evaporated it leaves solid matter in the form of concentric rings on the watch glass. These are deposits of dissolved salts left after evaporation of tap water.

Significance of dissolved salts in water

The significance of dissolved salts in water :

1. The dissolved salts provide moderate taste to water.

2. These salts are required for the growth of plants and various living organisms.

3. These salts provide various minerals required to our body for different biological functions.

Water contains dissolved air

Experiment 2 : To show water contains dissolved gases

Procedure :

Take a round bottom flask filled with water and fitted with a delivery tube using a cork. Put the other end of the delivery tube in an inverted graduated tube filled with water which is placed in a beaker. Heat water in the round bottom flask on the Bunsen burner.

Observation :

On heating the water in round bottom flask water boils and the gas bubbles are seen escaping from the flask. These bubbles pass through the delivery tube and are collected in the inverted graduated tube in the form of bubbles.

Solubility of gas is often expressed as volume of gas that can dissolve in a certain volume of solvent or solution at a specified temperature or pressure. Pure hydrogen gas, H_2, is almost purely water insoluble, and like all other gases, its solubility decreases as the temperature of the water increases. At 0°C, the solubility of hydrogen is 0.0019 g/kg of water, which drops to below 0.0012 g/kg at 60°C.

***Composition of dissolved air in water* :** Oxygen = 33%, Nitrogen = 66%, Carbon dioxide = 1%.

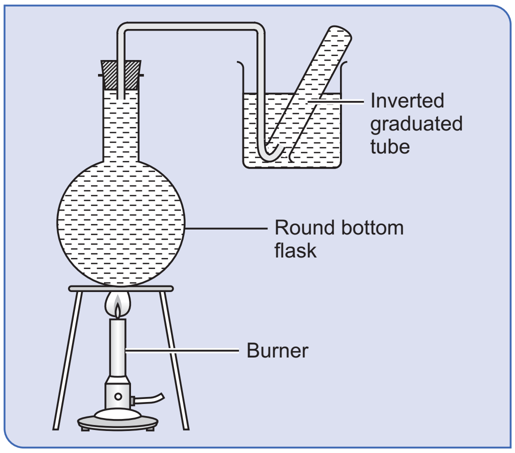


Fig. 3.3 : To show water contains dissolved gases

Significance of Air Dissolved In Water

The air dissolved in water has huge biological significance. Without air dissolved in water life in water would have been impossible. The key importance of air dissolved in water :

1. Dissolved oxygen is absolutely essential for the survival of all aquatic organisms

2. Carbon dioxide is present in water in the form of a dissolved gas. Aquatic plant life depends upon carbon dioxide in water for photosynthesis.

$$6CO_2 + 12H_2O \leftrightarrow C_6H_{12}O_6 + 6CO_2 + 6H_2O$$

Carbon dioxide also helps in conversion of limestone present in water bed to form calcium bicarbonate. Marine animals with shells (such as snails) convert this calcium bicarbonate to calcium carbonate to make their shells.

$$\underset{\text{(Limestone)}}{CaCO_3} + CO_2 + H_2O \rightarrow \underset{\text{(Calcium bicarbonate)}}{Ca(HCO_3)_2}$$

3.5. SOLUTIONS

Solutions as Mixture of Solids in Water

Solution is a homogeneous mixture composed of two or more substances. In solutions, a solute is a substance that is dissolved in another substance, known as a solvent.

$\therefore$ **Solution = Solute + Solvent**

The composition or concentration of the solution can be changed with gradual change in relative amounts of the components present in the solution.

A solute is the substance being dissolved in a mixture or solution. The solvent is the component of the solution that is responsible for dissolving the solutes. If the different components of a mixture exist in different physical state, then the final solution will remain in the same phase as that of the solvent.

For example : A saline water solution is made by dissolving table salt (NaCl) in water. Here the final solution is in the liquid phase, so water (the liquid) is the solvent and salt (a solid) is the solute in this case.

In an **aqueous** solution, the solute is dissolved in water (for example, salts or sugar in water). On the other hand, if the solvent is other than water it is known as **non-aqueous** solution (for example Sulphur in Carbon disulphide).

Types of solutions :

(a) Dilute solutions : A dilute solution has a low concentration of the solute compared to the solvent. Dilution is the process of decreasing the concentration of a solute in solution, usually simply by mixing with more solvent. To dilute a solution means to add more solvent without the addition of extra solute.

(b) Concentrated solutions : A concentrated solution is one in which there is a large amount of substance present in a mixture. It contains a relatively large amount of solute mixed in the solvent. To concentrate a solution either you have to increase the amount of solute or decrease the amount of solvent (through evaporation).

(c) Saturated solutions : A saturated solution is one in which any additional solute added to the solution is no longer dissolved. As the temperature of the solvent increases, solubility increases. Additionally, pressure and the nature of the solute and solvent affects the solubility.

(d) Unsaturated solution : An unsaturated solution is a solution, which contains less amount of solute than is required to saturate it at that temperature. In other words, when a solution can have solute added and dissolved, the solution is unsaturated. As the temperature of the solvent increases, solubility increases. Additionally, pressure and the nature of the solute and solvent affects solubility.

(e) Supersaturated solution : If the temperature of a saturated solution is increased then a little more amount of solute can be dissolved in it and if then the solution is cooled a bit, it will become a supersaturated solution. A supersaturated solution is a solution with more dissolved solute than the solvent would normally dissolve in its current conditions. Supersaturation condition is achieved by dissolving a solute in one set of conditions, then transferring it to other conditions without triggering any release of the solute.

Concentration of a solution

In chemistry, concentration is the abundance of a constituent divided by the total volume of a mixture. It can be expressed in several ways :

(a) Molarity : It is the moles of solute per litre of solution.

(b) Mole fraction : It is the ratio of the number of moles of solute to the total number of moles of substances present.

(c) Mass percent : Mass percent is the mass of solid solute present in grams percent in 100 grams of the solution.

$$\text{Concentration of solution} = \frac{\text{Mass of solute}}{\text{Mass of solution (Solute + Solvent)}} \times 100$$

As an example, if 20 grams of sugar is added to 80 grams of water, the concentration of sugar in the solution is

$$\frac{20}{20+80} \times 100 = \frac{20}{100} \times 100 = 20\%$$

(d) Volume Percent : Volume percent is the volume of the solute in millilitres present in 100 ml of a solution and is expressed as :

$$\text{Volume percent} = \frac{\text{Volume of solute}}{\text{Volume of solute + Volume of solvent}} \times 100$$

As an example, if 20ml of alcohol is mixed with 80ml of water to have a solution of 100 ml then the volume percent of alcohol in the solution is :

$$\frac{20}{20+80} \times 100 = \frac{20}{100} \times 100 = 20\%$$

EXAMPLES

Example 1. A solution contains 40 g of common salt dissolved in 280 g of water. Calculate the concentration in mass percentage of the solution.

Solution : Mass of common salt (solute) = 40 g

Mass of water (solvent) = 280 g

$\therefore$ Mass of solution = Mass of solute + Mass of solvent

$= 40 \text{ g} + 280 \text{ g}$

$= 320 \text{ g}$

$$\text{Mass percentage} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

$$= \frac{40 \text{ g}}{280 \text{ g}} \times 100$$

$$= \frac{400 \text{ g}}{28} = 14.28\%$$

Example 2. 30 ml of ethyl alcohol is mixed with 70 ml of petrol. Find the volume percentage of the solution.

Solution : Volume of ethyl alcohol (solute) = 30 ml

Volume of petrol (solvent) = 70 ml

$\therefore$ Volume of solution = Volume of solute + Volume of solvent

$= (30 + 70)$ ml

$= 100$ ml

Volume percentage

$$= \frac{\text{Volume of solute}}{\text{Volume of solution}} \times 100$$

$$= \frac{30}{100} \times 100$$

$$= 30\%$$

INTEXT QUESTIONS

1. 30 g of sodium chloride is dissolved in water to make up the volume of the solution to 240 ml. Calculate the mass by volume percentage of sodium chloride in the solution.
2. 40 g of a solute is dissolved in 360 g of water. What is the mass percentage of the solution?
3. A solution contains 30 g of urea in 120 g of solution. Calculate the mass percentage of the solution.
4. A solution has been prepared by mixing 8.5 ml of methanol with 90 ml of water. Calculate the volume percentage of the solution.
5. Calculate the mass percentage of the solution when 30 g of silver is dissolved in 140 g of water.

3.6. SOLUBILITY

Solubility is a chemical property of a substance, where the solute dissolves in a solvent. It is measured in terms of the maximum amount of solute dissolved in a solvent at equilibrium. The solubility of different substances is different in the same solvent.

A substance is called insoluble in a solvent when it has a negligible solubility in it. For example : silver chloride in water has solubility 0.000015g.

When the substance has solubility in a solvent more than negligible, but less than highly soluble then it is called sparingly soluble in that solvent. For example, calcium hydroxide in water has solubility 0.17g.

Determination of solubility of a solute at a particular temperature

To determine the solubility of a solute in a particular solvent at a specific temperature the following procedure is to be followed :

A clean dry evaporating dish is to be weighed and the temperature of the saturated solution is to be noted.

Let M = the weight of dry evaporating dish and M_1 = weight of the dish + Saturated solution

Now Heat the solution to dryness

Thus M_2 = Weight of the dried dish after heating the solution.

$\therefore$ Mass of saturated solution = $M_1 - M$

and mass of solute = $M_2 - M$

Thus the mass of solvent = Mass of saturated solution – Mass of solute

$$= (M_1 - M) - (M_2 - M)$$

So at that temperature :

$$\boxed{\text{Solubility} = \frac{\text{Mass of solute}}{\text{Mass of solvent}} \times 100}$$

Factors Affecting Solubility

The rate of solubility or dissolution of a solid in a liquid depends on several factors such as :

(a) Temperature : Solubility increases with rise in temperature and decreases with fall in temperature

(b) Nature of solute or solvent : Solubility of polar solute is higher in polar solvent and that of non-polar solute in non-polar solvent.

(c) Size of solute particles : Total surface area is greater if particle size is small. Thus, dissolution is faster.

(d) Stirring : Stirring increases the contact of solute particles with the solvent, resulting in increased rates of solution formation.

Solubility Curve

The solubility curve shows a relationship between the solubility and temperature. The solubility curve is the graphical representation of the solubility changes of a solute in a solution with change in temperature. The solubility is plotted along the *Y*-axis while the temperature is indicated along the *X*-axis.

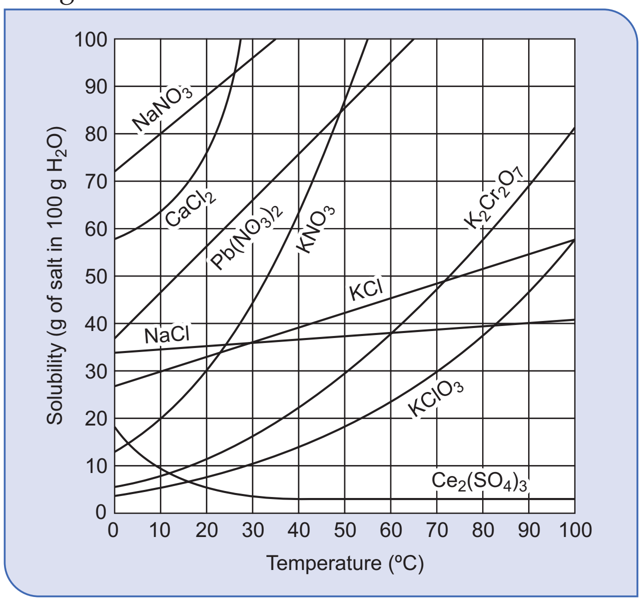


Fig. 3.4 : Solubility curve for common salts in water

From the solubility curve it is observed that :

(a) The solubility curve of calcium sulphate shows that, there is a decrease in solubility with further rise in temperature.

(b) The solubility of substances like sodium nitrate, potassium nitrate, and potassium bromide increases with rise in temperature.

(c) There is a little increase in the solubility of sodium chloride with increase in rise in temperature.

EXAMPLES

Example 1. 19 g of saturated solution of sodium chloride at 25°C, when evaporated to dryness, leaves a solid residue of 9 g. Calculate the solubility of sodium chloride.

Solution : Weight of water in solution 19 g – 9 g = 10 g

10 g of H_2O dissolves 9 g of solid

$\therefore$ 100 g of water will dissolve $\frac{9}{10} \times 100 = 90$ g

Thus, solubility of NaCl in H_2O at 25°C = 90 g

Example 2. Find the weight of sodium nitrate required to prepare 80 g pure crystals from its saturated solution at 80°C. Solubility of sodium nitrate is 150 g at 80°C and 100 g at 25°C.

Solution : Solubility at 80°C = 150 g

Solubility at 25°C = 100 g amount of crystals obtained when the solution is cooled from 80°C to 25°C

$$= 150 - 100 = 50 \text{ g}$$

To obtain 50 g of crystals, solium nitrate taken is 150 g.

To obtain 80 g crystals, sodium nitrate required will be $\frac{150}{50} \times 80 = 240$ g.

Anomalous Solubility

There are some salts whose solubility first increases with rise in temperature and then shows a fall after certain temperature such property of salts is known as **anomalous solubility.**

Example : Glauber salt ($Na_2SO_4.10\ H_2O$). The solubility curve of Glauber salt raises up to 32.8°C and then begins to fall a little as it is hydrous below the temperature 32.8°C and anhydrous above it.

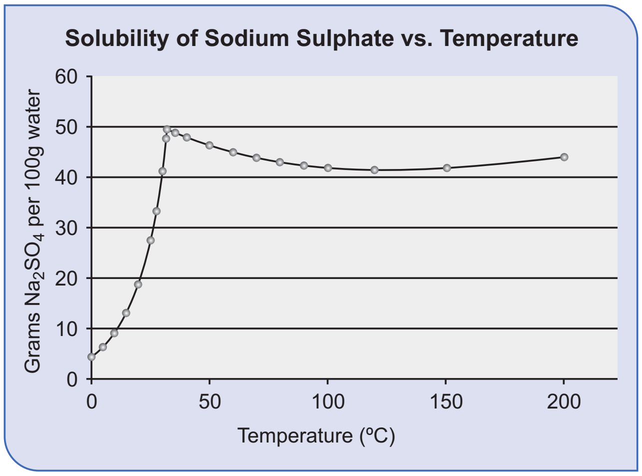


Fig. 3.5 : Solubility curve of glauber salt

Applications of Solubility Curves

(a) It is possible to determine the solubility of a substance at a particular temperature from the solubility curve.

(b) Solubility of different substances at different temperatures can be compared by the solubility curve.

(c) The effect of cooling of hot saturated solutions of different substances can be found from the curves.

(d) Knowledge of the solubility curve of different substances helps in the fractional crystallization of the substances.

(e) A discontinuity in the solubility curve indicates that the two different substances are involved.

Effect of pressure and temperature on solubility of gases in water

Pressure : An increase in pressure results in more molecules of the gas striking the surface of the liquid and entering the solution at a given time.

Therefore, at high pressure, more gas is dissolved in a given volume of liquid than at lower pressures.

This proportionality of the solubility of a gas to pressure was first proposed by William Henry, and is known as **Henry's law**.

Henry's law : Henry's law states that at a constant temperature, the amount of a gas that dissolves in a liquid is directly proportional to the partial pressure of that gas in equilibrium with that liquid, *i.e.*

$$m \propto p$$

$$m = Kp$$

where K = Proportionality constant
m = Mass of the gas
p = Partial pressure of the gas

Carbonated water (also known as soda water, sparkling water, or fizzy water) is water into which carbon dioxide gas is dissolved under high pressure. High pressure allows some more CO_2 to get dissolved in water and opening the soda water bottle suddenly decreases the pressure making the gas coming out as bubbles.

Temperature : The solubility of gas in water decreases with increase in temperature. On boiling, water loses its taste as the taste of water is due to the dissolved gases in it. If we boil water all the gases escape from the water making it tasteless.

3.7. HYDRATED SUBSTANCES

A hydrated substance is the crystalline substance that contains a definite number of water molecules. Specific numbers of water molecules are attached to each formula unit in a hydrated substance. For example : Washing soda having the formula $Na_2CO_3.10H_2O$ contains 10 molecules of water of crystallization.

Water of Crystallization

The water molecules that forms a part of the structure of a crystalline substance are called the water of crystallization. Water of crystallization is the part of the crystalline substance and water is often incorporated in the formation of crystals from the aqueous solutions. Many crystalline substances do not contain water of crystallization.

Table 1 : Examples of the substances having water of crystallization

Name	Chemical Formula	Common Name
Barium chloride	$BaCl_2.2H_2O$	—
Calcium chloride	$CaCl_2.6H_2O$	Dow flake
Calcium nitrate	$Ca(NO_3)_2.4H_2O$	Lime salt petre
Calcium sulphate	$CaSO_4.2H_2O$	Gypsum
Calcium sulphate semi hydrate	$(CaSO_4) . \frac{1}{2} H_2O$	Plaster of Paris
Copper(II) chloride	$CuCl_2.2H_2O$	—
Copper(II) sulphate	$CuSO_4.5H_2O$	Blue vitriol
Ferrous sulphate	$FeSO_4.7H_2O$	Green vitriol
Zinc sulphate	$ZnSO_4.7H_2O$	White vitriol
Magnesium sulphate	$MgSO_4.7H_2O$	Epsom salt
Potassium aluminium sulphate	$K_2SO_4.Al_2 (SO_4)_3.24H_2O$	Potash alum
Sodium carbonate deca hydrate	$Na_2CO_3.10H_2O$	Washing soda (crystal)
Sodium sulphate	$Na_2SO_4.10H_2O$	Glauber's salt

Experiment 3 : To demonstrate that the hydrated copper sulphate crystals (blue vitriol) contains water of crystallization.

Procedure :

Take small amount of powdered copper sulphate crystals in a clean and dry test tube. Clamp it in a tilted position to prevent condensed water to trickle back. Heat the crystals in that position above 100°C.

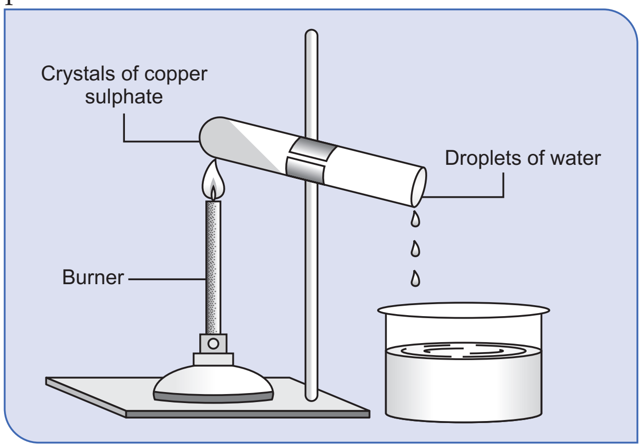


Fig. 3.6 : Effect of heat on hydrated copper sulphate

Observation :

Drops of colourless and odourless liquid is formed on the cooler part of the test tube. The residue left behind is anhydrous copper sulphate with amorphous (non-crystalline) structure (having no definite structure or shape).

The blue crystals of copper sulphate thus turned into white powder (non-crystalline) on heating in a test tube which turns back to blue crystals upon addition of a few drops of water.

$$\underset{\text{Blue crystals}}{CuSO_4.5H_2O} \xleftrightarrow{\text{Heat}} \underset{\text{Cool white powder}}{CuSO_4 + 5H_2O}$$

Crystalline substances without water of crystallization

- Sodium chloride (common or table salt) (NaCl)
- Nitre (KNO_3 = Potassium nitrate)
- Sugar ($C_{12}H_{22}O_{11}$)
- Potassium permanganate ($KMnO_4$)
- Ammonium chloride (NH_4Cl)

3.8. ANHYDROUS SUBSTANCES

A substance is anhydrous if it contains no water, for example, salts lacking their water of crystallization. The way of achieving the anhydrous form differs from one substance to another. Salts that contains water as part of their crystal structure are called hydrates (or hydrated salts) and the water in the crystal structure is called the water of hydration. When the water of hydration is removed from the hydrate, the salt that remains is said to be anhydrous. They are formed when the hydrated substances are heated at a temperature above 100°C they lose their water of crystallization and becomes amorphous.

3.8.1. Efflorescence

Efflorescence is the process where a compound loses its water of crystallization when exposed to dry air thus losing its crystalline shape and gets converted into powder. It is the loss of moisture of a salt to the surface of a porous material, where it forms a coating. Substances having efflorescence property are known as *efflorescent substances.*

When the vapour pressure in the hydrated substance is more than the atmospheric pressure then the phenomenon of the efflorescence takes place. **Examples of efflorescent substances** : Washing soda ($Na_2CO_3.10H_2O$); Glauber salt ($Na_2SO_4.10H_2O$).

3.8.2. Deliquescence

Deliquescence is the process, by which a substance absorbs moisture from the atmosphere, loses its crystalline form and ultimately dissolves in the absorbed water and forms a solution. Deliquescence occurs when the vapour pressure of the solution that is formed is less than the partial pressure of water vapour in the air. Substances showing deliquescent properties are called *deliquescent substances.* They readily absorb water out of the air.

Examples of deliquescent substances : Caustic soda (NaOH); caustic potash (KOH); magnesium chloride ($MgCl_2$); zinc chloride ($ZnCl_2$); calcium chloride ($CaCl_2$); ferric chloride ($FeCl_3$).

Common or table salt (NaCl) often seen to absorb moisture from the atmosphere, particularly in the rainy season and ultimately forms a solution. This is not because of the properties of sodium chloride (NaCl). In fact the commercial version of sodium chloride contains impurities such as magnesium chloride ($MgCl_2$) which are deliquescent substances. Pure sodium chloride has no deliquescent properties.

3.8.3. Hygroscopic substances

A hygroscopic substance is one that readily attracts water from its surroundings, through either absorption or adsorption when exposed to moist air. They are known as hygroscopic substances and the phenomenon is called as hygroscopy. Examples include honey, glycerin, ethanol, methanol, concentrated sulphuric acid, and concentrated sodium hydroxide.

3.9. DRYING AND DEHYDRATING AGENTS

Drying or desiccating agents are compounds that can absorb moisture readily from other substances without involving in any chemical reaction with them. They absorb water and are usually hygroscopic substances.

Anhydrous calcium or zinc chloride, phosphorus pentoxide, dry magnesium or sodium sulphate are some of the common examples of desiccating agents which absorbs water vapour or moisture from the air and are common desiccating or drying agents.

Usually most of the hygroscopic substances are desiccating agents such as concentrated sulphuric acid, phosphorus pentoxide, silica gel and quicklime.

Dehydrating agents are however substances that can even remove the chemically combined water molecules from the compounds. As an example, concentrated sulphuric acid can even remove the water molecule present in blue vitriol ($CuSO_4.5H_2O$).

Methods of drying gases

- By passing the moist gases through the columns of concentrated sulphuric acid
- By passing them through the drying tower (U-tube) filled with anhydrous sodium sulphate
- By passing the gases through drying bulb filled with anhydrous calcium chloride

Basic gases such as NH_3 can be dried by passing it through quicklime (basic in nature).

3.10. SOFT WATER AND HARD WATER

Hard water is water that has high mineral content compared to soft water. Hard water is formed when water percolates through deposits of limestone and chalk which are largely made up of calcium and magnesium carbonates.

Soft water is water that is free from dissolved salts of metals such as calcium, iron, or magnesium, which forms insoluble deposits such as soap curds in bathtubs and laundry equipment.

3.10.1. Cleansing Capacity of Soap with Hard and Soft Water

Although soap is a good cleaning agent, its cleaning capacity is reduced when used in hard water. Hardness of water is due to the presence of sulphates, chlorides or bicarbonate salts of Ca^{2+} or Mg^{2+}ions. Soaps are sodium or potassium salts of long chain fatty acids. When soap is added to hard water, the Ca^{2+} and Mg^{2+} ions present in hard water reacts with soap. The sodium salts present in soaps are converted to their corresponding calcium and magnesium salts which are precipitated as scum. The insoluble scum sticks on the clothes and so the cleaning capacity of soap is reduced.

$$Ca^{2+} + \underset{\text{Soap}}{RCOONa} \rightarrow \underset{\substack{\text{Calcium salt}\\ \text{(Scum)}}}{RCOOCa} + 2Na^{+}$$

$$Mg^{2+} + \underset{\text{Soap}}{RCOONa} \rightarrow \underset{\substack{\text{Magnesium salt}\\ \text{(Scum)}}}{RCOOMg} + 2Na^{+}$$

3.10.2. Advantages and disadvantages of hard water

Advantages of Hard Water

(1) Hard water is good for making our bones and teeth stronger when we drink it. The reason hard water is capable of doing this is because it contains high amounts of calcium and iron, which are good for the bones and teeth. This means the more hard water you drink, the stronger your bones and teeth are going to become.

(2) Hard water does not dissolves lead, and as a result of this it does not leads to lead poisoning if lead pipes are used in transferring the water into households.

(3) Hard water tastes better than soft water.

(4) Hard water is generally safer for drinking than soft water.

(5) Many industries such as Tanneries prefers hard water for curing leather.

Disadvantages of Hard Water

(1) Since hard water does not lather easily with soap, it wastes a great deal of soap when it is used in washing. It therefore is not economical to be used in washing.

(2) It is not advisable to use hard water in washing white fabrics since it tends to stain white fabrics by making them appear grey. More often than not when you use hard water to wash your white clothes, you are going to see the clothes turning grey after you have washed them. This is what hard water often does to white fabrics.

(3) Hard water is not good for dyeing materials. This is why the dyeing industry does not use it to work.

(4) Hard water forms annoying lime scales in containers such as kettles, pots, pipes, etc.

(5) It is harmful for boilers.

3.10.3. Types and Causes of Hardness

Calcium and magnesium dissolved in water are the two most common minerals that makes water "hard." Thus the degree of hardness becomes greater as the calcium and magnesium content increases and is related to the concentration of multivalent cations dissolved in the water.

The hardness of water is of two types :

(a) Temporary hardness

(b) Permanent hardness

(a) Temporary hardness : It is due to the presence of soluble hydrogencarbonates of calcium and magnesium in water and it can be easily removed by boiling.

(b) Permanent hardness : It is due to the presence of soluble chlorides and sulphates of calcium and magnesium in water and it cannot be removed by boiling.

3.10.4. Removal of Hardness

There are a number of methods to remove the hardness present in water. The common methods are :

(a) Boiling : Boiling the water can remove the temporary hardness of water or in other words the temporary hardness in water can be easily removed by boiling. On boiling, calcium or magnesium bicarbonate decomposes to give calcium or magnesium carbonate, which is insoluble in water. Therefore, it precipitates out.

$$\underset{\substack{\text{Calcium}\\ \text{bicarbonate}}}{Ca(HCO_3)_2} \xrightarrow{\text{Boiling}} \underset{\substack{\text{Calcium}\\ \text{carbonate}\\ \text{(insoluble)}}}{CaCO_3\downarrow} + H_2O + CO_2$$

(b) Adding Slaked Lime (Clark's Process) : Addition of slaked lime [$Ca(OH)_2$] can remove the temporary hardness of water as insoluble calcium carbonate precipitates out.

$$\underset{\text{Calcium bicarbonate}}{Ca(HCO_3)_2} + \underset{\text{Slaked lime}}{Ca(OH)_2} \xrightarrow{\text{Boiling}} \underset{\text{Calcium carbonate}}{2CaCO_3\downarrow} + 2H_2O$$

$$\underset{\text{Magnesium bicarbonate}}{Mg(HCO_3)_2} + \underset{\text{Slaked lime}}{Ca(OH)_2} \rightarrow \underset{\text{Magnesium carbonate (insoluble)}}{MgCO_3\downarrow} + CaCO_3\downarrow + 2H_2O$$

(c) By adding washing soda (Sodium carbonate): Boiling cannot remove the permanent hardness of water, one of the method which can be employed for the removal of hardness of water is the use of washing soda. Washing soda reacts with soluble calcium and magnesium chlorides and sulphates in hard water and forms the insoluble calcium carbonate and magnesium carbonate which gets precipitated.

$$\underset{\text{Calcium Sulphate}}{CaSO_4} + \underset{\text{Sodium Carbonate}}{Na_2CO_3} \rightarrow \underset{\text{Calcium Carbonate insoluble}}{CaCO_3\downarrow} + \underset{\text{Sodium Sulphate}}{Na_2SO_4}$$

$$\underset{\text{Magnesium Chloride}}{MgCl_2} + \underset{\text{Sodium Carbonate}}{Na_2CO_3} \rightarrow \underset{\text{Magnesium Carbonate insoluble}}{MgCO_3\downarrow} + \underset{\text{(Sodium Chloride)}}{2\,NaCl}$$

SUMMARY

- Water is made up of two elements, hydrogen and oxygen where two hydrogen atoms are bonded to a single oxygen atom to form water molecule. Its chemical formula is H_2O.
- The existence of water is essential for life on Earth.
- Water has three different states, solid, liquid and gas.
- The word water usually refers to water in its liquid state. The solid state of water is known as ice while the gaseous state of water is known as steam or water vapour.
- Water covers around 71% of the Earth's surface.
- Water from a sea or ocean is known as seawater. On average, every kilogram (2.2 lb) of seawater contains around 35 grams of dissolved salt.
- The freezing point of water lowers as the amount of salt dissolved in it increases. With average levels of salt, seawater freezes at – 2 °C (28.4 °F).
- Water is a good solvent and makes homogenous solution with sugar, salts and acids. On the other hand oils and fats do not mix with water. Water is known as the universal solvent.
- The water cycle involves water evaporation, rising to the sky, cooling and condensing into tiny drops of water or ice crystals that we see as clouds, falling back to Earth as rain, snow or hail before evaporating again and continuing the cycle.
- Pure water has no smell and no taste, it also has a pH level around 7.
- Water expands as it cools from 4°C to 0°C (above 4°C it does the opposite).

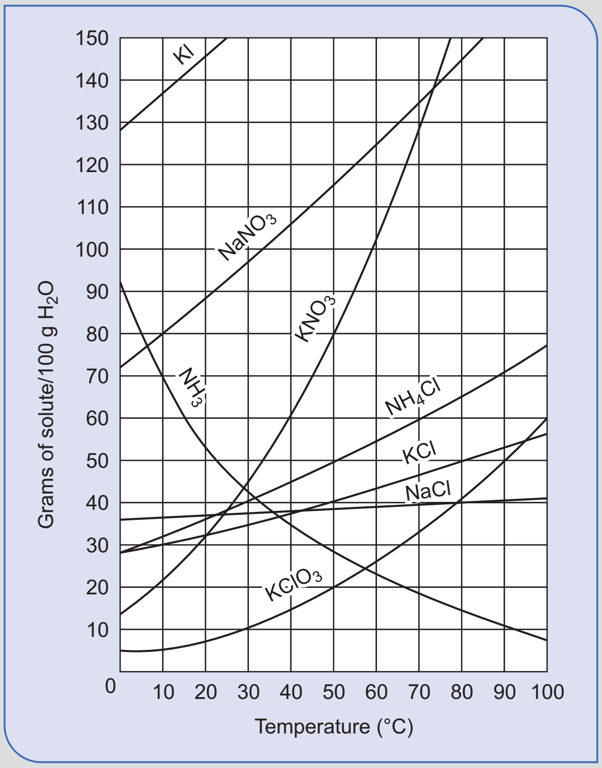
EXERCISE

A· VERY SHORT ANSWER TYPE QUESTIONS

1. Why water is known as the universal solvent?
2. Give one property which confirms water to be a compound?
3. Give the reaction for the formation of water molecule.
4. What is the Heat of vaporisation of water?
5. What is Latent heat of fusion?
6. What is a super-saturated solution?
7. What is the solubility of a compound?
8. What is water of crystallization?
9. Give the formula of the following :
   (a) Gypsum (b) Green vitriol
   (c) Blue vitriol (d) White vitriol
10. Give the two examples of the substance which have no water of crystallization.

B· SHORT ANSWER TYPE QUESTIONS

1. Why can't we rinse the soap off our hands with hard water ?
2. Why is water the universal solvent?
3. Explain why increasing the surface area of a solid solute helps to dissolve it more quickly.
4. Explain why the ocean is considered a solution.
5. List 3 types of solutions with examples of each.
6. Identify 2 ways that can be used to make a gas dissolve faster in a liquid.
7. What explains why solids become more soluble as temperature increases and why gases becomes less soluble ?



Consult the above solubility curve and answer the following questions.

(a) Which is more soluble $NaNO_3$ or KCl ?

(b) How does the line drawn for a particular substance relate to the saturation of a solution of that substance ?

(c) How many grams of NH_4Cl will dissolve in 100 grams of 90°C water ?

(d) If you are asked to make a saturated solution of KCl in 100 grams of water, what other piece of information would you need before you could start ?

(e) At what temperature will 10 grams of $KClO_3$ dissolve ?

(f) At what temperature will equal amounts of KNO_3 and NH_3 dissolve ?

(g) Which substance's solubility is LEAST affected by increasing temperatures ?

(h) Which substance's solubility is MOST affected by increasing temperatures?

(i) Which substance is most soluble at 0° Celsius ?

(j) If half as much water were used how would the amount of solute dissolved be affected for all of these substances ?

8. What is Henry's law ?
9. Define efflorescence with example.
10. Define deliquescence with example.
11. What is a hygroscopic substance ?
12. What do you mean by drying and de-hydrating agents ?
13. What do you mean by hydrated and an-hydrous substances ?
14. What is water of crystallization ?
15. What do you mean by soft and hard water ?

C· LONG ANSWER TYPE QUESTIONS

1. What is water cycle ?
2. How will you confirm water to be a compound ?
3. State three unique properties of water.
4. What is anomalous expansion of water ?
5. Why water is called the universal solvent ?
6. Design an experiment to confirm water has dissolved solids.
7. Write the significance of dissolved solids in water.
8. Design an experiment to confirm water has dissolved air.
9. Write the significance of dissolved air in water.
10. State the properties of a saturated solution.
11. What is concentration and volume percent of a solution ?

12. How will you determine the solubility of a solute at a particular temperature ?

13. State the factors that affects solubility.

14. Explain anomalous solubility with example.

15. Design an experiment to prove copper sulphate has water of crystallisation.

16. What are the advantages and disadvantages of soft water ?

17. What are the advantages and disadvantages of hard water ?

18. Write the different types and causes of hardness with methods of removal of hardness of water.

D· FILL IN THE BLANKS

1. Water is made up of two elements ______ and ______.

2. The solid state of water is known as _______.

3. When water is cooled it gets _______.

4. Water freezes at _______ °C temperature.

5. The chemical formula of water is ________.

6. Pure water has a pH level of around ________.

7. In a solution, the substance that is being dissolved is the _____.

8. The oceans are an example of _____ solution.

9. A solution that contains all the solute it can hold at a given temperature is _____.

10. Increasing the surface area of a solid _____ the rate of solution.

11. When a gas is dissolved in a liquid, the gas dissolves faster if the liquid is _____.

12. The concentration of a solution that contains much solute in the solvent could be described as _____.

13. An alloy is an example of a _____ solution.

14. Adding more solute to a solvent _____ its boiling point.

15. The amount of solute that can be dissolved in a specific amount of solvent at a given temperature is its _____.

16. The concentration of a solution, in terms of mass, depends on the mass of the _____.

[**Ans :** 1. hydrogen, oxygen, 2. ice, 3. expands, 4. 05. H_2O, 6. seven, 7. solute, 8. liquid, 9. saturated, 10. speeds up, 11. cooled, 12. concentrated 13. solid, 14. increases, 15. solubility, 16. solute.]

E· TRUE OR FALSE

1. Can the average human survive without water for a few days or a few weeks ?

2. Pure water is tasteless.

3. Water is an example of a chemical element.

4. Does water cover more than 50% of the Earth's surface ?

5. Water boils at 100°C (212°F).

6. Increasing the temperature always increases the solubility of a solute in a solvent.

7. The type of solution depends on the state of the solute.

8. Polluted water does not creates a huge problem to the land animals.

9. Plants can grow without water.

[**Ans :** 1. False 2. True 3. False 4. False 5. True 6. False 7. False 8. False 9. False]

F· MATCH THE COLUMN

	Column A	Column B
1.	A crystal of solute was dropped into a solution and it dissolved. The original solution was	(a) Increases
2.	If a crystal of solute is dropped into a solution and other crystals appear, the solution was	(b) Decreases
3.	The rate of dissolving a solid in a liquid upon stirring	(c) Supersaturated
4.	The rate of dissolving a gas in a liquid upon stirring	(d) Increases
5.	The rate of solution for a gas in a liquid on decreasing temperature	(e) Unsaturated

[**Ans :** 1. (e), 2. (c) 3. (d), 4. (b) 5. (a)]

G· MULTIPLE CHOICE QUESTIONS

1. How much of the earth's surface is covered by water?
(a) 75% (b) 50%
(c) 65% (d) 70%

2. What percentage of the world's water supply is fresh?
(a) 2.5% (b) 7.5%
(c) 10% (d) 20%

3. How many percent of the world water is suitable for human use?
(a) 25% (b) 10%
(c) less than 1% (d) 40%

4. Water that is safe to drink is called :
(a) Tap water (b) Potable water
(c) Distilled water (d) Piliated water

5. What are the main sources of water pollution in India?
(a) Municipal sewage
(b) Bathing in rivers
(c) Industrial discharge
(d) Both (a) and (c)

6. The circulation of water between ocean and land is a continuous process known as :
(a) Transpiration (b) Evaporation
(c) Raining (d) Water cycle

7. Which of the following is not a source of potable water :
(a) Well (b) River
(c) Lake (d) Sea

8. Whether we obtain water from tap or well, it originally came from :
(a) Sea (b) Ocean
(c) River (d) Rain

9. A mixture that has uniform composition throughout is called :
(a) Saturated solution
(b) Super saturated solution
(c) Solution
(d) Solvent

10. Solution of mercury with other metals is called :
(a) Amalgam
(b) Saturated solution
(c) Supersaturated solution
(d) Unsaturated solution

11. Solubility is amount of solute dissolved in :
(a) 50 g of solvent
(b) 100 g of solvent
(c) 250 g of solvent
(d) 200 g of solvent

12. At a particular temperature solution which cannot dissolve more solute is called :
(a) Saturated solution
(b) Unsaturated solution
(c) Aqueous solution
(d) Supersaturated solution

13. Fog is a solution of
(a) Air in water vapours
(b) Air in air
(c) Vapors in vapours
(d) Water vapours in air

14. Solid solutions of metals are called :
(a) Alloys
(b) Amalgams
(c) Saturated solutions
(d) Supersaturated solution

15. In solutions particles are
(a) Invisible
(b) Visible by naked eye
(c) Visible by ordinary microscope
(d) Visible by electron microscope

16. Concentration of a dilute solution :
(a) is high
(b) is in comparable
(c) can be both (a) and (b)
(d) is low

17. Substance present in lesser amount in solution is :
(a) Solute
(b) Solvent
(c) Aqueous solvent
(d) None of above

18. An unsaturated solution :
(a) cannot dissolve more solute
(b) cannot add more solvent
(c) can add more solute
(d) cannot be made

19. Concentration of solution is the :
(a) Quantity of solvent in solute
(b) Quantity of solute in given solvent
(c) Unit to measure concentration
(d) Quantity of gas dissolved in liquid

20. In presence of crystals of solute a supersaturated solution is :
(a) not stable
(b) stable
(c) cannot be made
(d) none of above

21. One of most common solvent used for crystallisation is :
(a) Water
(b) Syrup
(c) Normal saline
(d) Sulphuric acid

22. Water which forms scum with soap is called :
(a) Hard water
(b) Soft water
(c) Distilled water
(d) Undistilled water

23. Hardness which can be removed by boiling is called :
(a) Permanent hardness
(b) Temporary hardness
(c) Stiffness
(d) Toughness

24. Temporary hardness can be removed by adding :
(a) Lime (b) Carbon
(c) Oxygen (d) Slaked lime

25. Addition of washing soda removes :
(a) Softness of water
(b) Temporary hardness of water
(c) Permanent hardness of water
(d) Hydrogen from water

26. Temporary hardness is caused due to :
(a) Magnesium carbonate
(b) Calcium sulphate
(c) Magnesium sulphate
(d) Magnesium chloride

27. Permanent hardness is caused due to :
(a) Magnesium carbonate
(b) Magnesium bicarbonate
(c) Magnesium sulphate
(d) All of the above

28. Which of the following combinations would increase the solubility of a gas in liquid ?
(a) Decrease the pressure and decrease the temperature
(b) Decrease the pressure and increase the temperature
(c) Increase the pressure and decrease the temperature
(d) Increase the pressure and increase the temperature

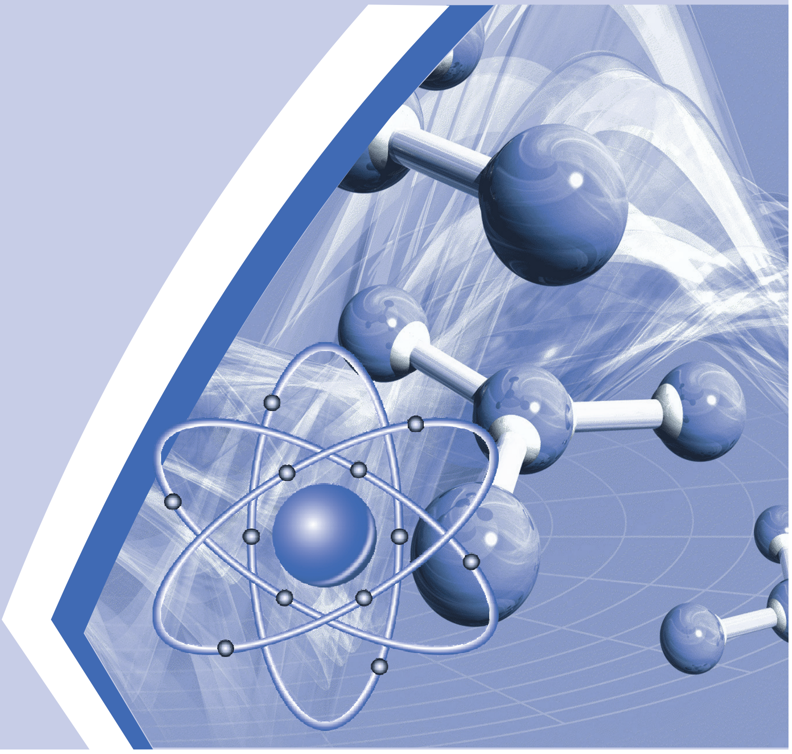
29. Which of the following will speed up the dissolving of a solid in water ?
(a) Cool the solution
(b) Freeze the solute
(c) Grind up the solvent
(d) Stir the solution

[**Ans :** 1. (d) 2. (a) 3. (c) 4. (b) 5. (d) 6. (d) 7. (d) 8. (b) 9. (c) 10. (a) 11. (b) 12. (a) 13. (d) 14. (a) 15. (a) 16. (d) 17. (a) 18. (c) 19. (b) 20. (a) 21. (a) 22. (a) 23. (a) 24. (d) 25. (c) 26. (a) 27. (c) 28. (c) 29. (d)]

✦ ✦ ✦

CHAPTER 4

ATOMIC STRUCTURE AND CHEMICAL BONDING



LEARNING OUTCOMES

4.1. Introduction
4.2. Sub-Atomic Particles of Atom
4.3. Structure of Atom
4.4. Electron Distribution in Orbits
4.5. Chemical Bonding

4.1. INTRODUCTION

We all know that the smallest particle of the substance is an atom. Atoms and molecules are the fundamental units of matter. The atom of one element is different from other elements. Many scientists came up with theories to put forth the structure of atom. The Greek philosopher **Democritus** coined the word 'atom' which means indivisible in Greek. Later **John Dalton** put forth the first scientific theory on the structure of atom. His theory had following postulates :

- Matter is made up of indivisible and small (cannot be seen by naked eye) particles called atom. Atoms are held together by forces of attraction.
- Atoms can neither be created nor be destroyed.
- Two or more atoms combine to form molecules of elements of compounds.
- Atoms of a given element have same properties and structure.
- Atoms are the smallest units that can undergo a chemical reaction.

However, these postulates did not hold good for isotopes. The forces of attraction between atoms are now referred to as chemical bonds. Later, it was discovered that atoms are not indivisible but made up of sub-atomic particles. In 1897, Cathode rays were discovered by **J. J. Thomson** and other scientists through a series of experiments. He proposed that the cathode rays consisted of negatively charged particles, named as electrons. **E. Goldstein** in 1886 discovered Anode rays also called canal rays. In 1932, **James Chadwick** proved the presence of heavy, neutral particles through a series of scattering experiments.

Name of Particle	Mass	Charge	Symbol
Neutrons	Mass of 1 H atom	No	${}_{0}^{1}n$
Proton	Mass of 1 H atom	Unit positive	${}_{+1}^{1}p$
Electron	1/1837 times mass of 1 H atom	Unit Negative	${}_{1}^{0}e$

4.2. SUB-ATOMIC PARTICLES OF ATOM

Atom is defined as the smallest particle of the element that can exist independently and retain all its chemical properties. Through a series of experiments, it was concluded that electrons, protons and neutrons are sub-atomic particles present in an atom.

Electrons : Electron is a sub-atomic particle which carries one unit negative charge and has a mass equal to 1/1837th of that of hydrogen. It is denoted by e. Electrons were discovered by **J. J. Thomson** in 1897. Electrons are negatively charged particles and are component of cathode rays and β-particle. Charge to mass ratio (e/m) of electrons is 1.76×10^8 C/g and does not depend on nature of the gas.

Protons : Proton is a sub-atomic particle which carries one unit positive charge and has a mass equal to that of an atom of hydrogen. It is denoted by '*p*'. Proton was discovered by **Goldstein**. These are positively charged particles and are component of anode rays. Charge to mass ratio (e/m) of proton is 9.58×10^4 C/g and depends on nature of gas.

Neutrons : Neutron was discovered by **James Chadwick** in 1932. It is denoted by 'n'. It is neutral in nature.

In the next topic, we will learn about the arrangement of these sub-atomic particles in an atom.

4.3. STRUCTURE OF ATOM

The discovery of electrons and protons led to the conclusion that atom is divisible and has an inner structure. Many theories were put forward to explain the arrangement of electrons and protons in an atom. J. J. Thomson was the first who proposed a model for the structure of atom.

4.3.1 Thomson Model of an Atom : *Discovery of Electrons*

J. J. Thomson proposed his model of the atom in 1904. Thomson model was compared to a watermelon. His theory had following postulates :

(i) An atom consists of a positively charged sphere with electrons filled in it.

(ii) The atom is electrically neutral because the positive and negative charges in an atom are equal in magnitude. It has no charge.

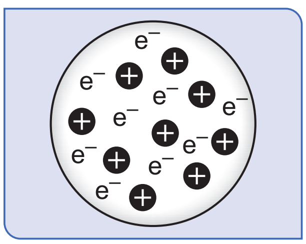


Fig. 4.1 : Thomson Plum Pudding Model

Thomson's model explained the electrical neutrality of an atom but it could not explain the results of other experiments performed at that time. Thus, it was rejected. The main drawbacks of the model were :

(a) This model failed to explain how the positive charges hold negatively charged electrons in an atom. Hence, it failed to explain the stability of an atom.

(b) It could not explain the position of nucleus of an atom.

(c) The results of the Scattering experiment and other experiment done at that time were not explained by this model.

Later J. J. Thomson studied the characteristics of cathode rays and conducted experiments on it using discharge tube. A discharge tube is connected to two electrodes. One to the positive terminal (anode) while another to the negative terminal (cathode). Vacuum is created inside the tube using a vacuum pump.

Properties of cathode rays :

- They are beams of negatively charged particles produced by the negative electrode, or cathode, in an evacuated tube.
- They travel towards the anode.
- They travel in straight lines and cast sharp shadows.
- They have energy.
- They are deflected by electric and magnetic fields and have a negative charge.

J. J. Thomson further concluded that :

- Cathode rays consist of negatively charged particles.
- Each negative charge has a definite mass and charge which remains constant for all elements.

Properties of electrons :

- They are integral part of an atom. Its mass is equal to 1/1837 of hydrogen atom.
- It carries a unit negative charge of -1.602×10^{-19} coulomb.

Goldstein Experiment : *Discovery of protons*

In his experiment, Goldstein used a discharge tube with a perforated cathode and air at low pressure around 0.001 mm Hg. At high voltage around 10,000 volt, he observed a faint red glow behind the cathode. The rays were formed at the anode and as stroked at the walls of the discharge tube, produced a faint red light. Goldstein named these rays as anode rays because they produced from anode. Since they produced from anode therefore, they have positive charge and attracted towards cathode. He concluded that anode rays move in straight lines because they cast a shadow of the objects which are placed in their way.

Properties of anode rays :

- Anode rays created a sharp shadow showing that they move in a straight line.
- Anode rays were deflected towards the negative plate showing that particles are positively charged.
- Anode rays were found to contain positively charged particle which was named as the proton.

Properties of proton :

- Proton carries a unit positive charge of 1.602×10^{-19} coulomb.
- Its mass is equal to 1837 times of an electron which is equal to 1.672×10^{-24}g.

4.3.2. Rutherford's Model of an Atom : *Discovery of Nucleus*

Rutherford tested Thomson's hypothesis by devising his "gold foil" experiment. Rutherford reasoned that if Thomson's model was correct then the mass of the atom was spread out throughout the atom. Rutherford's atomic model, also called nuclear atom or planetary model of the atom, description of the structure of atoms proposed (1911) by the New Zealand-born physicist Ernest Rutherford. The model described the atom as a tiny, dense, positively charged core called a nucleus, in which nearly all the mass is concentrated, around which the light, negative constituent, called electrons, circulate at some distance.

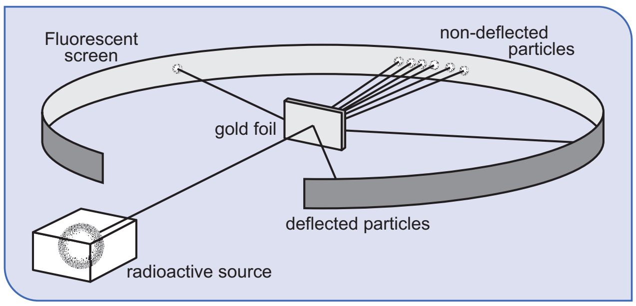


Fig. 4.2 : Gold foil experiment

In 1911, Rutherford performed scattering experiment by bombardment of thin foils of metals like gold, silver, platinum etc. with a beam of fast moving α particle. In his experiments he bombarded a thin sheet of gold foil with α-particles. He placed fluorescent zinc sulphide screen around the thin gold foil in order to study the deflection caused by the fast moving α-particles. He gave out following observations.

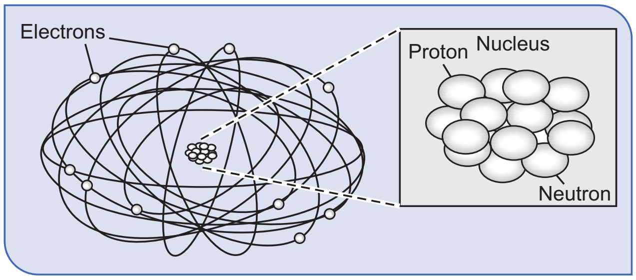


Fig. 4.3 : Rutherford's atomic model

Following observations were made by his experiment :

1. Most of the alpha particles pass through the foil straight without any deflection from their original path. This means that most of the space in an atom is empty.
2. A few alpha particles are deflected through small angles and a few are deflected through large angles. This shows that positive charge in an atom is not uniformly distributed.
3. A very few alpha particles completely rebound on hitting the gold foil and turn back on their path. This shows that volume occupied by the positively charged particles in an atom is very small as compared to the total volume of an atom. This model was compared to the solar system.

On the basis of his observations, Rutherford gave the nuclear model as :

(i) The positive charge and most of the mass of an atom was densely concentrated in extremely small region. This region was called as nucleus.

(ii) The magnitude of positive charge on the nucleus is different for different atoms.

(iii) The nucleus is surrounded by negatively charged electrons which balances the positive charge on the nucleus.

(iv) The electrons are not stationary but are revolving around the nucleus at a very high speed in circular path known as orbits.

(v) Electrons and nucleus are held together by electrostatic force of attraction.

(vi) Most of the space in an atom between the nucleus and revolving electrons is empty.

Drawbacks of Rutherford's Model

Rutherford's model was not able explain the stability of the atom. According to him, negatively charged electrons revolve around the nucleus in circular paths. When an object is moving in circular path, its motion is accelerated. Thus motion of electrons is accelerated. This means that when electrons move in circular path with accelerated motion, it will lose energy continuously by radiation. In this way, energy of revolving electrons will decrease gradually and therefore, their speed will go on decreasing. As a result, the electrons will be attracted more by nucleus and finally electrons should fall into the nucleus by taking a spiral path. This makes the atom unstable. But atoms are very stable. This is not explained by Rutherford.

Secondly, he did not tell anything about the arrangement of electrons in an atom.

4.3.3. Bohr's Model of Atom

In 1913, **Neil Bohr** proposed his quantized shell model of the atom to explain how electrons can have stable orbits around the nucleus. He was a Danish physicist and he got Nobel Prize in 1922 in physics for atomic structure and quantum mechanics. In order to explain the stability of an atom, Neils Bohr gave a new arrangement of electrons in the atom.

The main postulates of Bohr's model are :

(i) There are three fundamental particles in an atom : electrons, protons and neutrons. Electrons are negatively charged, protons are positively charged and neutrons have no charge. But the atom as a whole is electrically neutral.

(ii) Both the protons and neutrons are located in a small nucleus situated at the centre of the atom. Due to the presence of protons, the nucleus is positively charged.

(iii) The electrons are not found in the nucleus instead revolve rapidly round the nucleus in fixed circular paths called energy levels or shells. These can be represented as 1, 2, 3, 4, 5, 6 or K, L, M, N, O, P. These are counted from the centre and when an electron attains the lowest energy level it is said to be in a ground state.

(iv) The electrons in an atom move from the lower energy level to the higher energy level by gaining desired energy, and they move from higher energy level to the lower energy level by losing energy.

(v) Each energy level have a limited number of electrons. First energy level can hold maximum of 2 electrons, second shell can hold maximum of 8 electrons, third shell can hold maximum of 18 electrons and fourth shell 32 electrons.

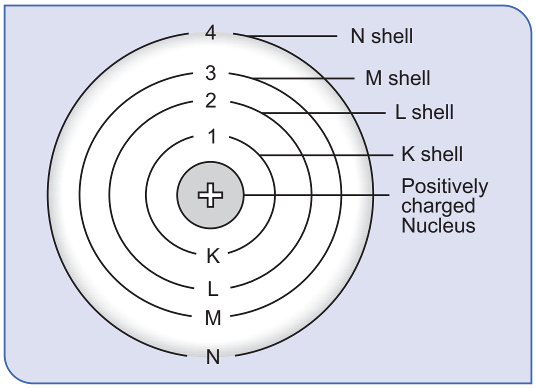


Fig. 4.4 : Energy levels in an Atom

(vi) With each energy level, a fixed amount of energy is associated. The shell nearest to nucleus has minimum energy and shell farthest from the nucleus has the maximum energy.

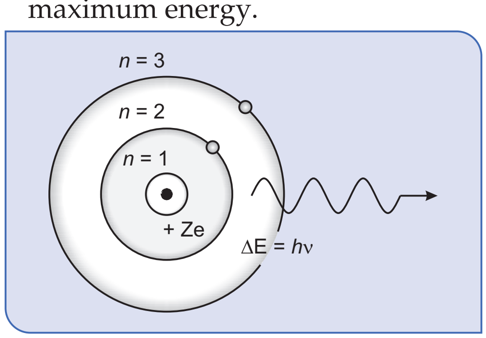


Fig. 4.5 : Energy change in an Atom

(vii) As long as the electron keep revolving in the same energy level, its energy is not changed. When an electron jumps from a lower energy level to a higher energy level, energy is absorbed. When an electron jumps from a higher energy level to lower energy level, energy is released. This explains the stability of atom.

(viii) The energy change is accompanied by absorption of radiation energy of

$$\boxed{E = E_2 - E_1 = h\nu}$$

where, h is a constant called 'Planck's constant' and ν is the frequency of radiation absorbed or emitted.

The value of h is **6.626×10^{-34} Js.**

INTEXT QUESTIONS

1. What is charge to mass ratio of electrons?
2. What are the sub-particles present in an atom ?
3. What name is given to negative particles of an atom ?
4. On what basis Thomson model was rejected ?
5. Write the distribution of electron in Sodium according to Bohr's model.

4.3.4. Atomic Number and Mass Number

Atomic number was determined by **Henry Moseley**. The number of protons present in nucleus of an atom is known as atomic number of that element. Atomic number is denoted by the letter 'Z'. No two elements have same atomic number. Atomic number can be used to identify an element. It does not change during a chemical reaction. Atomic number is equal to number of electrons only in case of neutral atom but not in ions. For example, Atomic number of nitrogen is 7 *i.e.*, Z =7. That means it contains 7 protons. Thus,

Atomic no. = Number of protons in nucleus

or **Z = Number of protons in nucleus of atom**

The total number of protons and neutrons present in the nucleus of an atom of an element is known as mass number. Mass number is denoted by letter A. Protons and neutrons are collectively called nucleons.

Mass number (A) = Number of protons (p) + Number of neutrons (n)

For example, Mass number of carbon is 12. It means it contains 6 protons and 6 neutrons. Another example is nitrogen whose, mass number is 14, it contains 7 electrons and 7 neutrons. Thus, mass number of an element is equal to sum of number of protons and neutrons.

Mass Number

Symbol of element

Atomic Number

or $^{A}_{Z}X$

where, A = Mass number

Z = Atomic number

X = Symbol of the element

4.4. ELECTRON DISTRIBUTION IN ORBITS

The distribution or arrangement of electrons in various shells of an atom of an element is called electronic configuration. To write the electronic configuration of an element, number of electrons in an atom of element and maximum number of electrons that can be accommodated in different shells should be known. Number of electrons will be known by atomic number of an element while the maximum number of electrons that can be accommodated in different shells was given by **Bohr and Bury**.

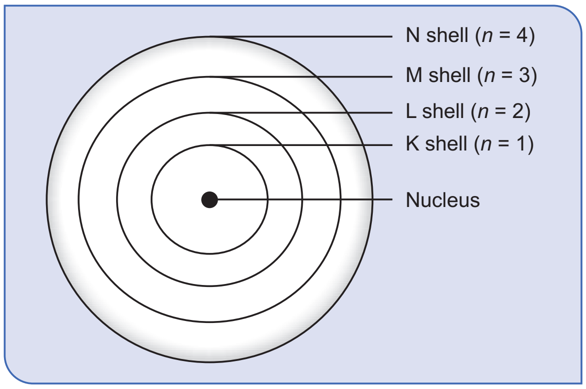


Fig. 4.6 : Distribution of electron in an atom

According to Bohr Bury scheme

(1) The maximum number of electrons which can be accommodated in any energy level of the atom is given by $2n^2$(where n is number of that energy level).

Example : For 1st energy level (K), $n = 1$

Maximum number of electrons = $2n^2 = 2(1)^2 = 2$

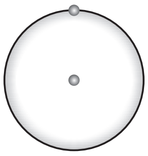
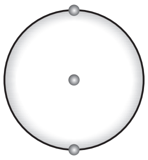
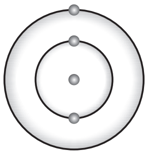
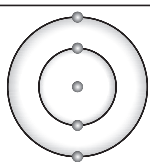
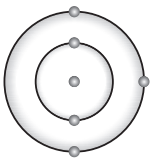
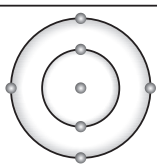
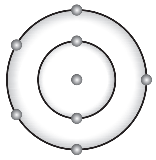
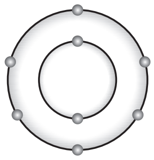
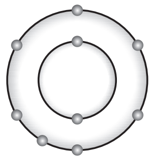
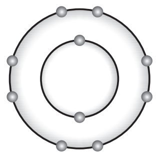
Similarly, second shell (L shell) contains maximum of 8 electrons, third shell (M shell) contains maximum of 18 electrons and fourth shell (N shell) contains maximum of 32 electrons.

(2) The outermost shell of an atom cannot accommodate more than 8 electrons, even if it has the capacity to accommodate more electrons.

Example : If N is the outermost shell of an atom, it can hold a maximum of eight electrons though its maximum rated capacity is 32 electrons. Eight electrons in valence shell gives stability.

(3) Electrons in an atom do not occupy a new shell unless all the inner shells are completely filled.

Electron distribution in first 10 elements is given below :

1. Hydrogen	
2. Helium	
3. Lithium	
4. Beryllium	
5. Boron	
6. Carbon	
7. Nitrogen	
8. Oxygen	
9. Fluorine	
10. Neon	

4.4.1. Valency

The combining capacity of an element is known as valency. It depends upon the number of electrons present in valence shell of an atom . Valence shell is the outermost shell of an atom. The electrons present in valence shell are known as valence electrons. In a chemical reaction, only the valence electrons take part. The valency of an element is either equal to the number of valence electrons in its atom or equal to the number of electrons required to complete eight electrons in the valence shell.

In case of metal, Valency is equal to the number of valence electrons. For example: Magnesium has to valence electrons and its valency is 2. In case of non-metals, valency is equal to 8 minus the number of valence electrons. For example, Nitrogen has five valence electrons. Therefore, its valency will be 3.

Valency may be of two types :

(a) **Electrovalency :** The number of electron lost or gained by one atom of an element to achieve nearest noble gas configuration is called electrovalency.

(b) **Covalency :** The number of electrons shared by an atom of an element to achieve nearest noble gas configuration is called covalency.

According to "Electronic theory of chemical bonding" of chemical combination, atoms can combine either by transfer of valence electrons from one atom to another (gaining or losing) or by sharing of valence electrons in order to have an octet in their valence shells and become stable. This is known as "**Octet Rule**". All the atoms combine to form compounds in such a way that each atom gets 8 electrons in its outermost shell. Atoms which have less than 8 electrons in the outermost shell are unstable. However, in case of Hydrogen, it can gain stability by having two electrons in the valence shell. This condition is known as duplet. Thus, the reason of chemical combination is to complete its octet and duplet in case of helium. The number of electrons gained or lost gives us directly the combining capacity of the element or valency of that element.

Table 1: Electron distribution and valency of first 20 elements :

At. No.	Element Name	Symbol	Electronic distribution				Valency
			K	L	M	N	
1.	Hydrogen	H	1				1
2.	Helium	He	2				0
3.	Lithium	Li	2	1			1
4.	Beryllium	Be	2	2			2
5.	Boron	B	2	3			3
6.	Carbon	C	2	4			4
7.	Nitrogen	N	2	5			3
8.	Oxygen	O	2	6			2
9.	Fluorine	F	2	7			1
10.	Neon	Ne	2	8			0
11.	Sodium	Na	2	8	1		1
12.	Magnesium	Mg	2	8	2		2
13.	Aluminium	Al	2	8	3		3
14.	Silicon	Si	2	8	4		4
15.	Phosphorous	P	2	8	5		3
16.	Sulphur	S	2	8	6		2
17.	Chlorine	Cl	2	8	7		1
18.	Argon	Ar	2	8	8		0
19.	Potassium	K	2	8	8	1	1
20.	Calcium	Ca	2	8	8	2	2

4.3.2. Isotopes

Any of two or more forms of a chemical element, having the same number of protons in the nucleus, or the same atomic number, but having different numbers of neutrons in the nucleus, or different atomic weights is called isotope. The word 'isotope' comes from the Greek 'isos' (meaning 'same') and 'topes' (meaning 'place') because the elements occupies the same place in the periodic table irrespective of being different in subatomic constitution. The atoms of some elements which have same atomic number but different atomic mass. For example, hydrogen has three different atomic species namely protium ($_1^1H$), deuterium ($_1^2H$) and tritium ($_1^3H$). The atomic number of each hydrogen is one but the mass number is 1, 2 and

3, respectively. Other such examples are carbon, $_6^{12}C$ and $_6^{14}C$, and chlorine, $_{17}^{35}Cl$ and $_{17}^{37}Cl$ etc. Isotopes are atoms of same element having same atomic number but different mass numbers. The chemical properties of isotopes are similar but their physical properties are different.

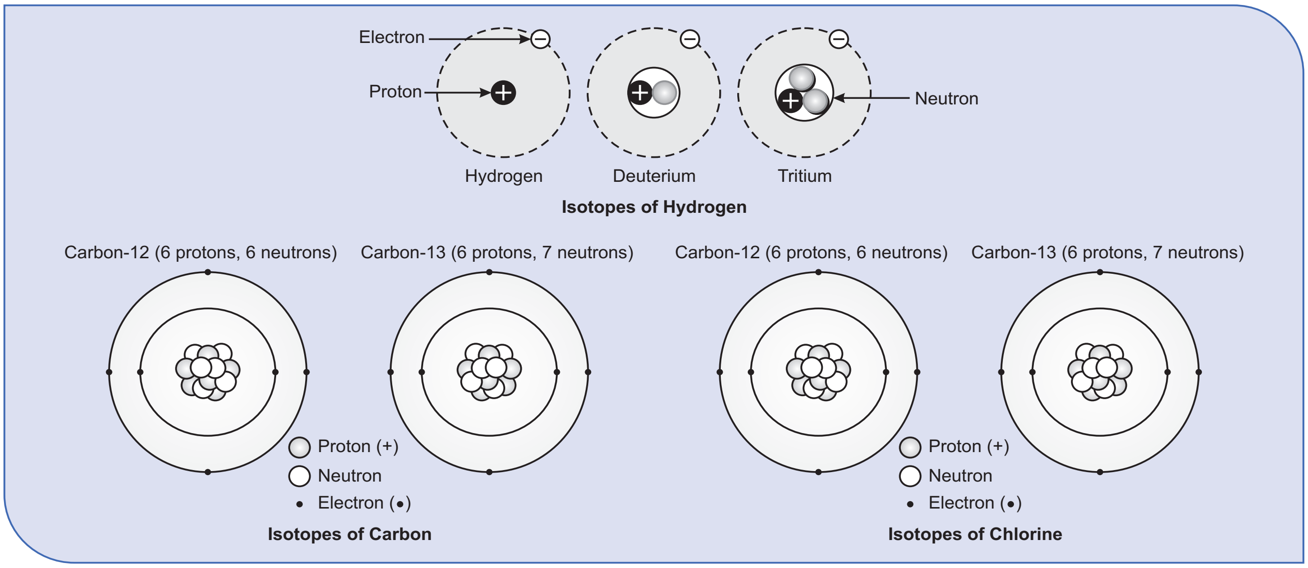


Fig 4.7 : Isotopes of Hydrogen, Carbon and Chlorine

In case of isotopes, the mass of an atom of any natural element is taken as the average mass of all the naturally occurring atoms of that element. If an element has no isotopes, then the mass of its atom would be equal to as the sum of protons and neutrons present in it. But if an element occurs in isotopic forms, then the average mass can be calculated by the percentage of each isotopic form. The average atomic mass of an element is the sum of the masses of its isotopes, each multiplied by its natural abundance.

Average atomic mass = $f_1M_1 + f_2M_2 + ... + f_nM_n$
where f is the **fraction** representing the natural abundance of the isotope and M is the mass number (weight) of the isotope.

Applications of Isotopes

Some of the application of isotopes are :

(i) Treatment of cancer : Radioactive isotope cobalt-60 is used to cure cancer.

(ii) Fuel in nuclear reactors : Uranium-235 is used as a fuel in nuclear reactors for generating electricity. This process is known as nuclear fission i.e., breaking of heavy nucleus into two smaller nuclei with the liberation of large amount of energy.

(iii) To detect leakage in underground pipes : Radioactive isotopes are used in industry to detect the leakage in underground oil pipelines, gas pipelines and water pipes.

(iv) As tracers to detect tumours : Radioactive isotopes are used to detect the presence of tumor and blood clots in human body.

INTEXT QUESTIONS

1. If, in atom, K shell is filled and L shell has 4 electron. What will be the atomic number of that element ?
2. Calculate the valency of Boron and Oxygen.
3. If the number of electron in a neutral atom is 11. Then what will be the valency of that element ?
4. Calculate number of protons, electrons and neutrons is $_{35}Br^{80}$.
5. Given that the percentage abundance of the isotope ^{35}Cl is 75% and that of ^{37}Cl is 25%. Calculate the average atomic mass of chlorine.
6. What are the number of neutrons in isotopes if carbon $_6^{12}C$ and $_6^{13}C$.
7. Is there any change in electronic configuration of isotopes ?

4.5. CHEMICAL BONDING

As we all know that except noble gases, all other elements of the Periodic table have incomplete octet. They combine with other elements to complete their octet and acquire stable inert gas configuration. In this process, they either transfer or share one or more electrons with other elements and forms a bridge between them. This force of attraction which holds atoms together in the molecule of a compound is called a chemical bond.

Chemical bond may be of three types :

(a) Ionic or electrovalent bond
(b) Covalent bond
(c) Coordinate covalent bond

4.5.1. Ionic or Electrovalent Bond

The chemical bond is formed by the complete transfer of valence electrons from one atom to another is called electrovalent bond or ionic bond. The atom which loses electrons forms positive ions called cation and the other species which accepts electrons forms negative ions called anion. This electrostatic attraction between positive and negative ions is called ionic bond.

Characteristics of ionic compounds :

(i) Ionic compounds are good conductors of electricity in solution or in molten state not in solid state due to presence of ions and arrangement of ions in crystal lattice.

(ii) Ionic compounds have generally high melting and boiling points due to strong electrostatic force of attraction. Due to this, large amount of energy is required to separate them.

(iii) Ionic compounds are not directional and do not show resonance.

(iv) Ionic reactions are fast and simple. For example, neutralization, precipitation etc.

(v) These compounds are generally soluble in polar solvents such as water, liquid ammonia etc. due to its ions but insoluble in non-polar solvents like CCl_4, CS_2, etc.

Examples of Ionic Compounds

(1) Formation of sodium chloride (NaCl) : The electronic configuration of sodium (Z= 11) is 2, 8, 1. It loses one electron from its outermost shell and attains the noble gas configuration of neon.

$$\underset{2,8,1}{Na} \longrightarrow \underset{2,8}{Na^+} + e^-$$

Similarly the electronic configuration of Chlorine (At. no. 17) is 2, 8, 7. It accepts the electron donated by outer shell of sodium and attains the stable configuration of noble gas Argon.

$$\underset{2,8,7}{Cl + e^-} \longrightarrow \underset{2,8,8}{Cl^-}$$

When sodium and chlorine combines with each other, ionic bond is formed as a result of electrostatic force of attraction between them.

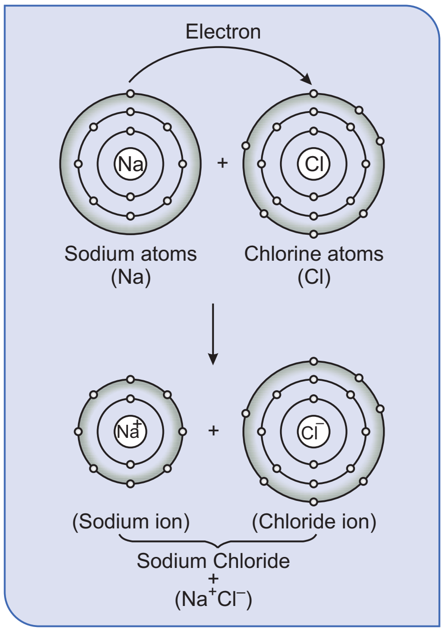


Fig. 4.8 : Formation of Sodium Chloride Molecule

(2) Formation of Magnesium Chloride ($MgCl_2$) : Magnesium (At. No. 12) has electronic configuration 2, 8, 2. It donates outermost two electrons and forms positive ion by attaining the configuration of nearest noble gas Argon.

$$\underset{2,8,2}{Mg} \longrightarrow \underset{2,8}{Mg^{2+}} + 2e^-$$

On the other hand, the electronic configuration of chlorine (At. No. 17) is 2,8,7. Hence, each chlorine atom accepts one electron donated by magnesium and forms two negative ions. Each chloride ions have one negative charge.

$$\underset{2,7}{2Cl + 2e^-} \longrightarrow \underset{2,8}{2Cl^-}$$

These oppositely charged ions attracts each other by electrostatic force of attraction and forms $MgCl_2$ molecule.

The **electrovalence** is thus equal to the number of unit charge(s) on the ion. Thus, magnesium is assigned a positive electrovalence of two, while chlorine a negative electrovalence of one.

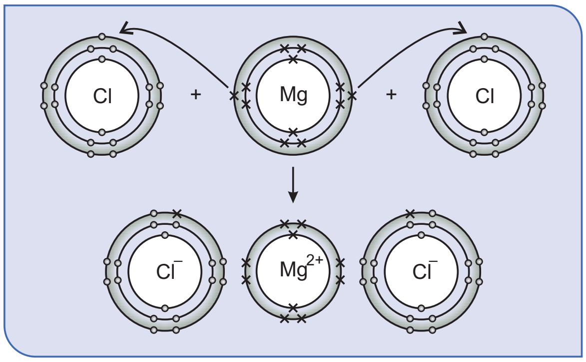


Fig. 4.9 : Formation of Magnesium Chloride molecule

(3) Formation of Calcium Oxide (CaO) : Calcium (At. No. 20) has electronic configuration 2,8,8,2. It donates outermost two electrons and forms positive ion by attaining the configuration of nearest noble gas Argon.

$$\underset{2,8,8,2}{Ca} \longrightarrow \underset{2,8,8}{Ca^{2+}} + 2e^-$$

On the other hand, the electronic configuration of Oxygen (At. No. 8) is 2, 6. It accepts the electrons donated by Calcium and forms negative ion.

$$\underset{(2,\,6)}{O} + 2e^- \longrightarrow \underset{2,8,8}{O^{2-}}$$

These oppositely charged ions attracts each other by electrostatic force of attraction and forms CaO molecule.

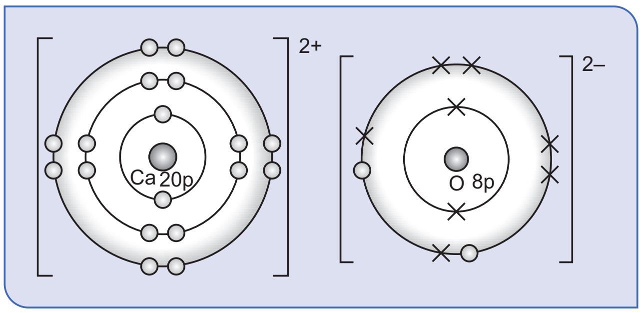


Fig. 4.10 : Formation of calcium chloride molecule.

4.5.2. Covalent Bond

Covalent bond was first proposed by G. N. Lewis in 1916. Covalent bond is formed by the mutual sharing of electrons between the combining atoms of same or different elements to complete its octet or duplet (hydrogen). Number of electrons shared by each atom is known as covalency. Covalent bond is represented by single short line (-).

Conditions for formation of covalent bond :

(1) Combining atom should be short by 1, 2 or 3 electrons in valence shell.

(2) Electronegativity difference should be very small or zero between the two atoms.

(3) Two atoms should approach one another accompanied by decrease of energy.

Characteristics of covalent compounds :

(a) Physical state : These compounds are generally liquids or gases or soft solids. **Exception :** ***diamond (Hardest substance).***

(b) Melting and boiling point : These compounds have generally low melting and boiling point. **Exception :** ***diamond, silica (high melting and boiling point).***

(c) Electrical conductivity : These compounds are bad conductors of electricity. **Exception :** ***graphite, which is a good conductor.***

(d) Solubility : These compounds are generally insoluble in polar solvents but soluble in non-polar solvents like benzene. **Exception :** ***alcohol, dissolves in water due to hydrogen bonding.***

(e) Molecular reactions : Covalent compounds show molecular reactions which are slow.

Types of Covalent bonds :

(i) Single covalent bond : When a covalent bond is formed by sharing of one electron from each atom, then it is called single covalent bond and denoted by (–) single short line. **For example :** Cl – Cl, H – H, H – Br etc.

(a) Formation of hydrogen molecule (H_2) : In the outermost shell of hydrogen atom, there is a one electron, hydrogen shares this electron with another hydrogen atom to acquire nearest inert gas configuration of helium. Thus, a single bond is formed between these two hydrogen atoms.

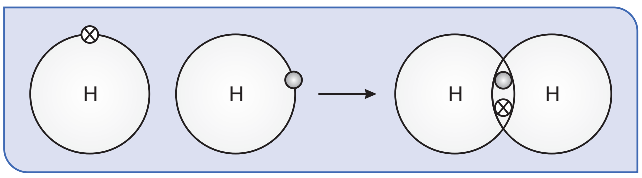


Fig. 4.11 : Formation of hydrogen molecule

(b) Formation of chlorine molecule (Cl_2) : In the outermost shell of chlorine atom, it has seven electrons, it share its one valence electron with another chlorine atom to form chlorine molecule by single covalent bond.

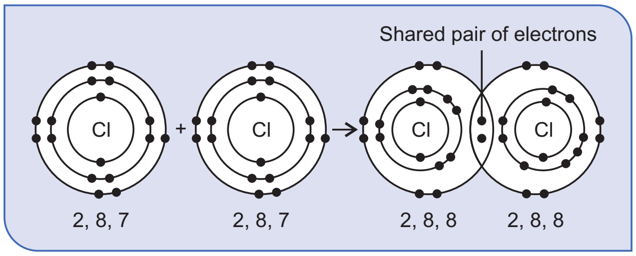


Fig. 4.12 : Formation of chlorine molecule

(ii) Double covalent bond : In a covalent bond, if two electrons are shared between two same or different atoms then it is called double covalent bond and denoted by two short lines (=). For example : O = O

Formation of oxygen molecule (O_2) : Oxygen atom has six electrons in its outermost shell of, it shares its two valence electrons with another oxygen atom to form oxygen molecule by double covalent bond. It is represented as O = O.

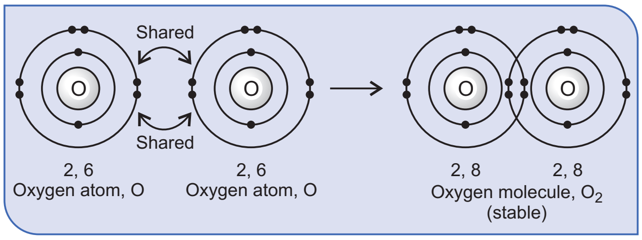


Fig. 4.13 : Formation of oxygen molecule

(iii) Triple Covalent Bond : When a covalent bond is formed by sharing of three electrons between each atom then it is called triple covalent bond, and denoted by three short lines (≡). For example : N≡N.

(a) Formation of Nitrogen molecule (N_2) : Nitrogen atom has five electrons in its outermost shell, it shares three valance electrons with another nitrogen molecule by triple covalent Bond. It is represented as N ≡ N.

The bond distance of multiple bonds are shorter and the bond energies are higher.

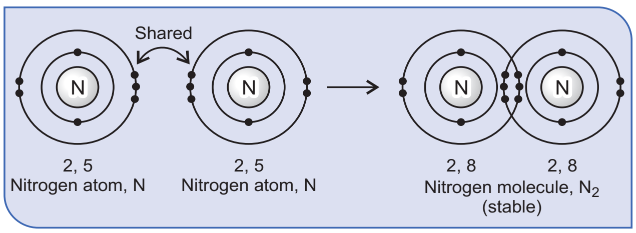


Fig. 4.14 : Formation of nitrogen molecule

Other examples of covalent bond :

(i) Hydrogen chloride : In this molecule, both hydrogen and chlorine share one electron each and complete their octet and duplet (hydrogen).

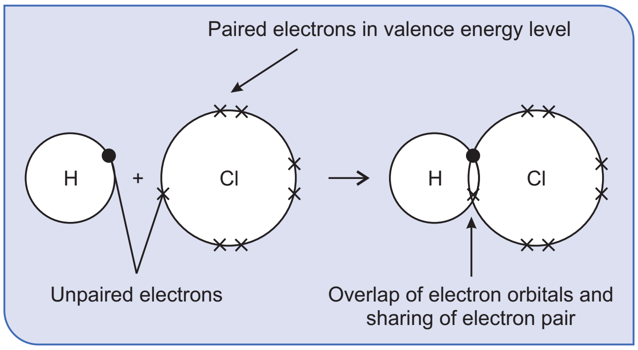


Fig. 4.15 : Formation of hydrogen chloride molecule

(ii) Carbon Tetrachloride : In this molecule four single bonds are formed between carbon and four chlorine atoms.

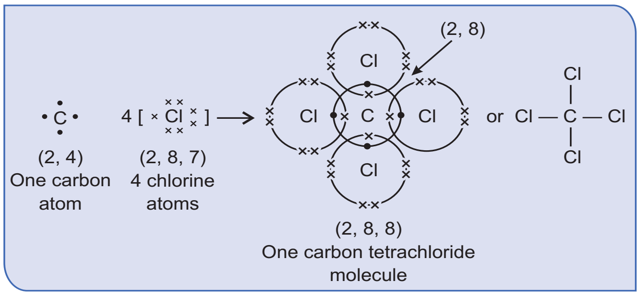


Fig. 4.16 : Formation of carbon tetrachloride molecule

(iii) Ammonia molecule : Hydrogen has one valence electron and nitrogen has five valence electrons. All the three atoms of hydrogen share their electrons with nitrogen and compete their duplet and nitrogen completes its octet while Hydrogen completes its duplet.

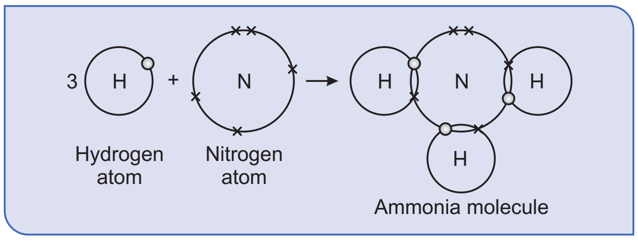
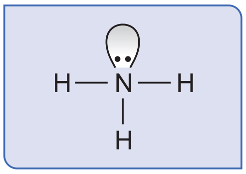


Fig. 4.17 : Formation of ammonia molecule

or



(iv) Methane molecule : Carbon has 4 valence electrons. To achieve stability, it should have 8 electrons in its valence shell. Hydrogen has 1 electron in its valence shell and it should have 2 electrons in its valence shell to achieve stability.

Thus, C shares each of its electrons with one hydrogen atom covalently linked to form CH_4 and achieve a stable form.

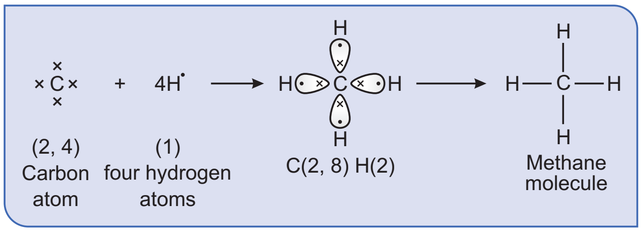


Fig. 4.18. Formation of Methane molecule

INTEXT QUESTIONS

1. Why in formation of ionic bond, metals should have low ionisation enthalpy and non-metals have strong electron gain enthalpy ?
2. Why covalent compounds are not soluble in water ?
3. Show the formation of double covalent bond with an example.
4. Write the electronic distribution of element having atomic number 12. State whether it is a metal or non-metal ? Draw its atomic structure diagram.
5. What type of compound is magnesium chloride ? Is it soluble in water or not ?
6. Which type of bond is formed in ammonia molecule ? Show its molecular structure.

SUMMARY

- An atom is defined as the smallest particle of the element that can exist independently and retains all its chemical properties.
- Electrons, protons and neutrons are sub-atomic particles present in an atom.
- J. J. Thomson proposed his model of the atom in 1904. In Thomson's model, the atom is composed of electrons which are surrounded by a soup of positive charge and balanced by electrons which carries the negative charge. This is also known as Plum Pudding Model.
- In 1911, Rutherford performed scattering experiment by bombardment of thin foils of metals like gold, silver, platinum etc. with a beam of fast moving α-particle. According to his model, the positive charge and most of the mass of atom was densely concentrated in the nucleus and electrons moves around the nucleus in circular paths called orbits.
- In 1913, Bohr proposed his quantized shell model of the atom. According to him, there are three fundamental particles in an atom : electrons, protons and neutrons. Electrons are negatively charged, protons are positively charged and neutrons have no charge. But the atom as a whole is electrically neutral. Electrons are present in circular paths or level and each energy level is associated with fixed amount of energy.
- The number of protons present in an atom is called atomic number.
- The total number of protons and neutrons present in the nucleus of an atom of an element is known as mass number.
- Electronic distribution was given by Bohr and Bury. It is known as Bohr Bury Scheme. According to it, number of electrons a shell can accommodate are 2, 8, 18 respectively.
- The combining capacity of an element is known as valency.
- All the atoms combine to form compounds in such a way that each atom gets 8 electrons in its outermost shell and this is Octet Rule.
- Isotopes are atoms of same element having same atomic number but different mass numbers.

- Chemical bond may be of three types :
  - (a) **Ionic or Electrovalent bond :** This bond is formed by transfer of electrons.
  - (b) **Covalent bond :** Covalent bond is formed by sharing of electrons.
  - (c) **Coordinate covalent bond :** It is a special type of covalent bond in which the shared pair of electrons comes from one species only but shared by both the atoms so as to complete their octets.

EXERCISE

A. VERY SHORT ANSWER TYPE QUESTIONS

1. Write the names of the particles which constitute an atom.
2. Who discovered protons ?
3. What is the valency of element having atomic number 15 ?
4. How many electrons at the maximum can be present in the first shell ?
5. What is an atom ?
6. What will be the valency of an element, the formula of whose oxide is A_2B_3 ?
7. What is atomic mass ?
8. Where is neutron located in an atom.
9. Name the atom which contains only one neutron in its nucleus ?
10. How many shells are occupied in oxygen atom ?
11. What is a chemical bond ?
12. What kind of elements form positively charged ions ?
13. Name the covalent compound, which dissolves in water and conducts electricity.
14. What type of bonding is found in fluoride molecule ?
15. Name two compounds in which covalent bond is present.
16. Name the electrons which are involved in bond formation between the atoms ?
17. Give the absolute mass of neutron.

B. SHORT ANSWER TYPE QUESTIONS

1. Write the drawbacks of Rutherford Model.
2. Define atomic number and mass number.
3. From the symbol $^{35}Cl_{17}$ state :
   (i) Atomic number of chlorine (ii) Mass number of chlorine (iii) Electronic configuration of chlorine
4. What is the atomic number, mass number and valency of an element having 3 protons, 3 electrons and 4 neutrons?
5. Why the valency of Argon is zero ?
6. Draw atomic structure for all the isotopes of hydrogen with number of protons and number of neutrons.
7. Write the symbol for the atom with the given atomic number [Z] and mass number [A].
   (i) Z = 8, A = 16 (ii) Z = 5, A = 10 (iii) Z = 16; A = 33.
8. Write the distribution of electrons in sodium and fluorine atoms.
9. Calculate the number of electrons, protons and neutrons in $^{35}Cl_{17}$, $^{40}Cl_{18}$ and $^{23}Cl_{11}$.
10. What are the conditions for the formation of covalent bonds ?
11. How electronegativity, electron affinity and ionisation potential affect formation of ionic bond ?
12. Explain the formation of sodium chloride molecule with diagram.
13. Why ionic bonds are formed between metals and non-metals ?
14. Why melting points of ionic compound are high ?
15. Why ionic compounds are soluble in water ?
16. Explain covalent bond with example.
17. State the types of bonds present in :
    (a) Sodium chloride
    (b) Methane
    (c) Carbon tetrachloride.
    (d) Calcium oxide
    (e) Ammonium ion
    (f) Oxygen molecule

C. LONG ANSWER TYPE QUESTIONS

1. Describe Thomson Model of an atom.
2. Give reason why ?
   (a) Solid NaCl does not conduct electricity but its aqueous solution conducts electricity.
   (b) Ionic compounds generally have high melting and boiling points.

(c) Ionic compounds are soluble in polar solvents but insoluble in non-polar solvents.
(d) Covalent compounds do not conduct electricity.
(e) Covalent compounds usually have low melting and boiling points.

3. Explain Bohr's model of an atom.
4. Write the characteristics of ionic compounds.
5. Explain types of covalent bonds with examples.
6. The electronic configuration of atoms of three elements is X = (2, 8, 1) Y = (2, 8, 6) and Z = (2, 8, 7).
(i) Write the formula of X and its atomic structure diagram by mentioning type of bond in it.
(ii) What will be the formula of compound formed between X and Z.
(iii) Classify the above elements as metals and non-metals.
(iv) Which element is a good conductor of electricity and why ?

7. Describe Bohr Bury Scheme for distribution of electrons in an atom.
8. (i) HCl is a covalent compound but it is soluble in water. Why ?
(ii) Explain the formation of nitrogen molecule with atomic structure.

9. Write the postulates of Rutherford's scattering experiment.

D· FILL IN THE BLANKS

1. …………radioisotope is used to cure cancer.
2. The isotope of hydrogen having mass number 3 is ………….
3. The covalency of nitrogen is ……..
4. Cathode rays are made of ………..charged particles.
5. ………..discovered nucleus.
6. The charge on neutrons is ……….
7. Covalent compounds are generally ……….. in water.
8. In chlorine molecule, there is …….. ..covalent bond between two chlorine atoms.
9. Melting and boiling point of ionic compounds is ……..
10. If the atomic number of an element is 12 then its atom contains 12………
11. The ground state configuration of an element E is 2, 8. The forces exist between its molecules is ……..
12. The maximum number of covalent bonds formed between molecules could not be more than ……..

[**Ans :** (1) Cobalt-60 (2) tritium (3) three (4) negatively (5) Rutherford (6) zero (7) insoluble (8) single (9) high (10) protons (11) vander Waals Forces (12)three]

E· TRUE OR FALSE

1. Isotopes of an element differ in the number of electrons in it.
2. Electrovalency and covalency of an element are numerically the same.
3. Since the atoms of inert gases are very stable or unreactive they cannot exist in the Free State.
4. The maximum number of electrons which can be accommodated in M shell is 32.
5. Mass number of any atom may have fractional value.
6. The e/m ratio of positive rays is maximum when hydrogen is used in the discharge tube.
7. Covalent bond is formed by transfer of electrons.
8. There is one ionic, three covalent and one coordinate covalent bond in Ammonium chloride.
9. The covalency of carbon is one.
10. In the molecule of HCl, covalent bonding is found.

[**Ans :** (1) False (2) True (3) False (4) False (5) False (6) True (7) False (8) True (9) False (10) True.]

F· MATCH THE COLUMN

Column A	Column B
1. J. J. Thomson	(a) Ionic bond
2. Sodium chloride	(b) Covalent bond
3. Moseley	(c) Valency 2
4. Methane	(d) Discovery of electrons
5. Magnesium	(e) Neutron
6. James Chadwick	(f) Atomic number

[**Ans :** 1. (d), 2. (a), 3. (f), 4. (b), 5. (c), 6. (e)]

(b)

Column A	Column B
1. Helium	(a) 2, 4
2. Nitrogen	(b) 2, 8, 8
3. Carbon	(c) 2
4. Argon	(d) 2, 8, 8, 2
5. Magnesium	(e) 2, 5
6. Calcium	(f) 2, 8, 2

[Ans : 1. (c), 2. (e), 3. (a), 4. (b), 5. (f), 6. (d)]

G· MULTIPLE CHOICE QUESTIONS

1. Electrons were discovered by :
(a) James Chadwick (b) J. J. Thomson
(c) E. Goldstein (d) None of these

2. Which is the lightest gas ?
(a) Argon (b) Neon
(c) Hydrogen (d) Kryptom

3. The charge on the neutron is :
(a) 1.6×10^{-19} coulomb (b) 1.6×10^{-24} gram
(c) 9×10^{-28} gram (d) 0

4. When chlorine accepts electrons from other elements, it acquires the configuration of :
(a) Argon gas (b) Helium gas
(c) Neon gas (d) Xenon gas

5. A bond between a metal and a non-metal is most likely to be :
(a) Ionic bond (b) Covalent bond
(c) Both (a) and (b) (d) Dative bond

6. What will be the maximum number of electrons in M shell ?
(a) 8 (b) 2
(c) 18 (d) 32

7. The reactions of ionic compounds are generally :
(a) Slow (b) Fast
(c) Can be fast or slow (d) None of these

8. For the formation of ionic bond, electro-negativity difference between the combining atoms should be :
(a) Equal to 1.7 eV (b) More than 1.7 eV
(c) Less than 1.7 eV (d) Both (a) and (b)

9. The particle not present in ordinary hydrogen atom is :
(a) Proton (b) Electron
(c) Nucleus (d) Neutron

10. The first model of an atom is given by :
(a) Neils Bohr
(b) Ernest Rutherford
(c) J.J.Thomson
(d) Eugen Goldstein

11. NaCl molecule has :
(a) Two covalent bonds (b) Ionic bond
(c) Three covalent bonds (d) None of these

12. The forces present between ionic compounds are :
(a) Electrostatic forces of attraction
(b) Vander Waal's forces
(c) Molecular forces
(d) None of these

13. The number of valence electrons in Na^+ ion is :
(a) 11 (b) 10
(c) 12 (d) None of these

14. The number of valence electrons in graphite atom is :
(a) 4 (b) 3
(c) 2 (d) 5

15. Number of electrons shared by each chlorine atom during formation of chlorine molecule is :
(a) 1 (b) 2
(c) 3 (d) 4

[Ans : 1. (b) 2. (c) 3. (d) 4. (a) 5. (a) 6. (c) 7. (b) 8. (d) 9. (d) 10. (c) 11. (b) 12. (a) 13. (b) 14. (a) 15. (a)]

✦✦✦

CHAPTER 5

THE PERIODIC TABLE

LEARNING OUTCOMES

5.1. Introduction
5.2. Classification of Elements
5.3. Discovery of Atomic Number and its Use in Modern Periodic Law
5.4. Modern Periodic Table
5.5. Special Reference to Alkali Metals, Alkaline earth Metals, Halogen and Zero group

5.1 INTRODUCTION

Earlier by 1865, only 31 elements were known out of 114 known elements. Gradually, with the discovery of elements, more information about their properties was gathered. Therefore, it is very difficult to study and remember each and every element and its properties. Hence, to study about such a large number of elements, they were classified according to their properties to make their study easier. The elements were classified in such a way that, the elements having similar properties were grouped together while, elements having dissimilar properties were separated from one another. This is known as classification of elements. "Periodic table is a tabular arrangement of the chemical elements according to their atomic number, electronic configurations and chemical properties". In olden days elements were classified into two main groups that is metals and non-metal based on the properties like appearance, rigidity, ductility, malleability etc. But this classification did not work well as some metal show metallic as well as non-metallic characteristics. In addition, there are lots of variations within metals and non-metals. Later on, many theories were developed and the concept of atomic mass came into existence.

Need for the classification of elements :

The systematic classification of the elements in the Periodic Table has the following advantages :

- It enables chemists to analyse and understand the properties of the elements and their compounds more systematically and orderly.
- It enables chemists to predict the properties of the elements and their compounds based on their positions in the Periodic Table, and vice versa.
- It becomes easier to study, understand, compare and contrast the related properties among the elements and their compounds from different groups.

5.2. CLASSIFICATION OF ELEMENTS

Many scientists came up with different ways and formats of arranging and classifying elements, to facilitate easy study and understanding.

5.2.1 : Dobereiner's Triad

In 1817, a German scientist **Johann Wolfgang Dobereiner,** showed the relation between properties of elements and their atomic masses. He divided elements having similar properties into groups of three and called them as triads. He observed that the atomic mass of the middle element of a triad was nearly equal to the arithmetic mean of atomic masses of other two elements. Thus, he put forth the Law of triads which states "When elements are arranged in order of the increasing atomic masses, the atomic mass of middle element is approximately equal to the arithmetic mean of the other two elements". The below table shows few examples of Dobereiner's Triad.

Alkali Triad		Atomic weight of Na
Metal	Atomic weight	$= \frac{39+6{\cdot}9}{2}$
Lithium	6.9 u	$= 22.95$ u
Sodium	23.0 u	
Potassium	39.0 u	
Alkaline Earth Triad		Atomic weight of Sr
Metal	Atomic weight	$= \frac{40{\cdot}1+137{\cdot}3}{2}$
Calcium	40.1 u	$= 88.7$ u
Strontium	87.6 u	
Barium	137.3 u	
Halogen Triad		Atomic weight of Br
Metal	Atomic weight	$= \frac{35{\cdot}5+126{\cdot}9}{2}$
Chlorine	35.5 u	$= 81.2$
Bromine	79.9 u	
Iodine	126.9 u	

In all the above examples, the atomic mass of middle element is average of atomic masses of other two elements of the triad.

Drawbacks of Dobereiner's Triad :

Some of the limitations of Dobereiner's Triad were :

(i) The law did not fully apply even within the same family. For instance, the three elements, nitrogen, phosphorus and arsenic have similar properties. Therefore, they can be arranged to form triad. But the actual atomic mass of the middle element phosphorus is 31 u, which is much lower than the arithmetic mean of the atomic masses of nitrogen and arsenic that is 44.45 u. Thus, these three elements do not fall into the Dobereiner's Triad in spite of having similar properties.

(ii) Dobereiner failed to arrange all the then known elements in the form of triads.

5.2.2. Newland's Law of Octaves

John Alexander Reina Newlands was an English chemist who invented periodic table. Newland divided the elements in horizontal rows of seven elements. His classification gave a very important conclusion that there is some systematic relationship between order of atomic masses and repetition of properties of elements. According to him,

"If elements be arranged in ascending order of their atomic masses, the properties of eighth element was repetition of properties of first element".The repetition resembles repetition of eighth note in an octave of music. So, it was known as **'Law of Octaves'**. For example,

Newlands arranged the elements then known in the following manner.

Newland's Periodic Table

H	F	Cl	Co/Ni	Br	Pd	I	Pt/Ir
Li	Na	K	Cu	Rb	Ag	Cs	Tl
G	Mg	Ca	Zn	Sr	Cd	Ba/V	Pb
Bo	Al	Cr	Y	Ce/La	U	Ta	Th
C	Si	Ti	In	Zn	Sn	W	Hg
N	P	Mn	As	Di/Mo	Sb	Nb	Bi
O	S	Fe	Se	Ro/Ru	Te	Au	Os

Limitations of Newland's Law of Octaves

(i) Law of Octaves is applicable to classification of higher elements having atomic mass up to 40 u, *i.e.* calcium only. After calcium every eighth element does not possess similar properties.

(ii) He assumed that there were 56 elements existed in nature and no new elements would be discovered in future. But later on, many new elements were discovered.

(iii) In order to fit the elements into his table, Newland not only put two elements together in same slot, but also placed some unlike elements in the same column. For instance, cobalt and nickel are placed in the same slot and these are placed in the same column as that of fluorine, chlorine and bromine which have very different properties.

INTEXT QUESTIONS

1. If the atomic mass of Calcium is 40 and Barium is 137, then what will be the atomic mass of Strontium if Ca, Sr and Ba forms a triad ?
2. Why classification according to Dobereiner was not accepted ?
3. What is the need of classificaton of elements ?
4. Which scientist gave the Law of Octaves? State the Law.
5. If any two elements obey Law of Octaves, how many elements are there between the two elements ?

5.2.3. Mendeleev Periodic Law

The first remarkable contribution to the classification of elements was given by a Russian Chemist, **Dmitri Ivanovich Mendeleev** (1834-1907), in 1869. Mendeleev started his work with only 63 elements which were known during that time. He arranged 63 elements on the basis of physical and chemical properties of elements, he gave a law known as the **Periodic law**. He arranged the elements in order of increasing atomic masses in horizontal rows in such a way that the elements with similar physical and chemical properties fell under the same vertical columns called groups.This law states that : ***When elements are arranged in order of their increasing atomic masses, elements with similar properties are repeated after certain regular intervals. This repetition of properties of elements after certain regular intervals is called periodicity of properties.***

Characteristics Features of Mendeleev Periodic Table

On the basis of this law, Mendeleev gave a table known as Mendeleev's periodic table.

(1) The horizontal rows present in the periodic table are called periods. There are seven periods in the periodic table and are numbered from 1 to 7 (Arabic numerals).

(2) Properties of elements in a particular period show regular gradation *i.e.* increase or decrease from left to right.

MENDELEEV PERIODIC TABLE

Groups	I		II		III		IV		V		VI		VII		VIII
Oxides Hydrides	RO RH		RO RH_2		R_2O_3 RH_3		RO_2 RH_4		R_2O_5 RH_3		RO_3 RH_2		R_2O_7 RH		RO_4
Periods ↓	A	B	A	B	A	B	A	B	A	B	A	B	A	B	Transition series
1	H 1.008														
2	Li 6.939		Be 9.102		B 10.81		C 12.011		N 14.007		O 15.999		F 18.998		
3	Na 22.99		Mg 24.31		Al 29.98		Si 28.09		P 30.974		S 32.06		Cl 35.453		
4 First series :	K 39.102		Ca 40.08			Sc 44.96		Ti 47.90		V 50.94		Cr 50.20		Mn 54.94	Fe 55.85 Co 58.93 Ni 58.71
Second series:		Cu 63.54		Zn 65.37	Ga 69.72		Ge 72.59		As 74.92		Se 78.96		Br 79.909		
5 First series :	Rb 85.47		Sr 87.62			Y 88.91		Zr 91.22		Nb 92.91		Mo 95.94		Te 99	Ru 101.07 Rh 102.91 Pd 106.4
Second series :		Ag 107.87		Cd 112.40	In 114.82		Sn 118.69		Sb 121.75		Te 127.60		I 126.90		
6 First series :	Cs 132.90		Ba 137.34			La 138.91		Hf 178.49		Ta 180.95		W 183.85			Os 190.2 Ir 192.2 Pt 195.09
Second series :		Au 196.97		Hg 200.59	Tl 204.37		Pb 207.19		Bi 208.98						

(3) Periods are nine in number and are designated as I, II, III, IV, V, VI VII, VIII and zero group.

(4) Except for zero and VIII group, all other groups are further divided into two subgroups designated as A and B.

(5) The elements on the left side of each group are designated as sub group A. These are called normal or representative elements.

(6) The groups on the right hand side of each group are designated as sub group B. These are known as transition elements.

(7) Group VIII contains nine transition elements in three sets each containing three elements. These three sets lie in IV, V and VI period.

(8) All the elements in a particular group are chemically similar in nature. They show regular gradation in their physical properties and chemical relativities.

Merits of Mendeleev Periodic Table

(1) Systemic study of elements : Mendeleev table simplified and systemize the study of the chemistry of the elements. He arranged known 63 elements into groups and periods. With the help of this table the study of chemical element has become simpler because, by knowing the properties of one element in a group, the properties of other elements can be easily predicted. Further while developing periodic table, he considered two parameters that is increasing atomic masses and grouping of similar elements together.

(2) Correction of atomic masses of elements : Mendeleev's periodic table helped to correct the doubtful atomic masses of some elements based on their positions in the periodic table. Similarly, atomic masses of gold, platinum, indium etc. were corrected on the basis of Mendeleev's Periodic table.

(3) Prediction of new elements and their properties : At the time of Mendeleev, only 63 elements were known while arranging these elements, he left some gaps. These gaps represent the elements which were not known at that time. Mendeleev predicted the properties of these undiscovered elements on the basis of their positions. For example scandium, gallium, and germanium were not known at that time. Mendeleev named these elements by prefixing eka meaning first to the name of the preceding element in the same group. That is eka-boron for scandium, eka-aluminium for gallium and eka-silicon for germanium which were discovered later. The observed properties of these elements were found to be similar to those predicted by Mendeleev.

(4) Positions of noble gases : During the predication of Mendeleev's periodic table, the noble gases like helium, neon, argon, krypton, xenon and radon were not known. After the discovery of these gases later in the year around 1900, they were placed in the gap separately in a new group called zero group without disturbing the existing order.

Limitations of Mendeleev Periodic Table

Despite so many advantages of Mendeleev periodic table, there are certain defects too. Some of them are :

(1) Position of hydrogen : Hydrogen is placed in group I, it resembles the elements of group I and group VII A. Thus, position of hydrogen is not correctly defined.

(2) Anomalous pairs of elements : Although the elements in Mendeleev's periodic table have been arranged in increasing order of their atomic masses, but in some cases, elements have been arranged on the basis of similarity in properties. As a result, an element having higher atomic mass has been placed before an element with lower atomic mass. **Example :** Cobalt (Co) with higher atomic mass (58.93 u) has been placed before nickel with atomic mass 58.71u.

(3) Position of isotopes : Isotopes are atoms of same element having same atomic number but different atomic masses. As Mendeleev's periodic table is based on atomic masses of elements, isotopes of an element have different atomic masses, therefore, they should be given different positions in the periodic table. Further, isotopes have same chemical properties therefore, they are placed in the same position in the periodic table. Thus, in the periodic table, the positions of the isotopes are anomalous.

(4) No resemblance of elements within sub-groups : In the Mendeleev periodic table, elements placed in a group having similar chemical properties. But there is no similarity in the properties of elements in the subgroups of a particular group. For instance : Li, Na, K, Rb and Cs present in group I A are quite different from coinage metals such as Cu, Ag, Au present in group IB.

(5) Uncertainty in prediction of new elements: In Mendeleev's periodic table, the elements are arranged in increasing order of their atomic masses. But the atomic masses do not increase in a regular pattern in going from one element to the next. Further, the difference in atomic masses between two successive elements is small among lighter elements but larger among heavier elements. Therefore, it is difficult to predict how many new elements could be discovered between two known elements especially among heavier elements.

(6) Different groups for similar elements : In Mendeleev's periodic table, some elements of similar properties are placed in different groups. For example, both copper and mercury have similar properties, but copper is placed in group I B, while mercury in II B.

INTEXT QUESTIONS

1. What is the basis of Mendeleev's Periodic Table ?
2. Write any two demerits of Mendeleev's Periodic table.
3. Mendeleev left gap for two elements and named them as eka-silicon and eka-aluminium. Name the elements that have taken place of these elements.

5.3. DISCOVERY OF ATOMIC NUMBER AND ITS USE IN MODERN PERIODIC LAW

Atomic number of an element is equal to number of protons present in an atom. In 1913, an English physicist, **Henry Mosley** exposed anodes made up of different metals in the discharge tube to cathode rays. Wavelength of rays of these anodes were different. Wavelength decreased in regular pattern. Thus, he concluded that positive charges present in the nucleus determines the identity and properties of the atom. The number of protons present in an atom is called as atomic number of that element.

The idea of atomic number formed the basis for discovery of Modern Periodic table. Modern periodic Law may be stated as :

"The physical and chemical properties of elements are periodic function of their atomic numbers".

Cause of periodicity : According to Modern Periodic Law, when elements are arranged in increasing order of their atomic numbers, the properties of elements are repeated after certain regular intervals. The cause of such periodicity in the properties of elements is the repetition of similar outer electronic configuration after certain regular intervals.

For example : All the elements of group I A *i.e.* alkali metals have similar outer electronic configuration ns^1, where n refers to number of outermost principal shell.

Table 1 : Electronic configuration of group IA elements

Element	Symbol	Atomic No.	Electronic Configuration
Lithium	Li	3	$[He]2s^1$
Sodium	Na	11	$[Ne]3s^1$
Potassium	K	19	$[Ar]4s^1$
Rubidium	Rb	37	$[Kr]5s^1$
Cesium	Cs	55	$[Xe]6s^1$
Francium	Fr	87	$[Rn]7s^1$

5.4 MODERN PERIODIC TABLE

On the basis of Modern Periodic Law, a table was proposed and is known **as Long form of Periodic Table or Modern Periodic Table** which is most commonly used. This table is based upon electronic configuration of elements.

In this table, elements are arranged in the order of their increasing atomic numbers. This table is also known as Bohr's table as it is based on Bohr's scheme for arrangement of various electrons around the nucleus that is electronic configuration of elements.

Periodic Table of the Elements

1 IA 1A	2 IIA 2A	3 IIIB 3B	4 IVB 4B	5 VB 5B	6 VIB 6B	7 VIIB 7B	8	9 VIII 8	10	11 IB 1B	12 IIB 2B	13 IIIA 3A	14 IVA 4A	15 VA 5A	16 VIA 6A	17 VIIA 7A	18 VIIIA 8A
1 **H** Hydrogen [1.00784;100811]																	2 **He** Helium 4.002602(2)
3 **Li** Lithium [6.938,6997]	4 **Be** Beryllium 9.012181(5)											5 **B** Boron [10806-10821]	6 **C** Carbon [12.0006;1201116]	7 **N** Nitrogen [14.00643;1400728]	8 **O** Oxygen [15.99903;15,99977]	9 **F** Fluorine 18.998403163(8)	10 **Ne** Neon 20.1797(6)
11 **Na** Sodium 22.9887682	12 **Mg** Magnesium [24,304,24.307]											13 **Al** Aluminum 26.9815388(8)	14 **Si** Silicon [23.034;28.086]	15 **P** Phosphorus 30.973761998(5)	16 **S** Sulfur [32.059:32.076]	17 **Cl** Chlorine [35,446;35.457]	18 **Ar** Argon 39.948(1)
19 **K** Potassium 38,0983(1)	20 **Ca** Calcium 40.078(4)	21 **Sc** Scandium 44.855908(5)	22 **Ti** Titanium 47.867(1)	23 **V** Vanadium 50.9415(1)	24 **Cr** Chromium 51,8897(6)	25 **Mn** Manganese 54,838045(5)	26 **Fe** Iron 55.845(2)	27 **Co** Cobalt 58.833194(4)	28 **Ni** Nickel 58.6234(4)	29 **Cu** Copper 83.546(3)	30 **Zn** Zinc 65.38(2)	31 **Ga** Gallium 68.723(1)	32 **Ge** Germanium 72.630(8)	33 **As** Arsenic 74.827595(6)	34 **Se** Selenium 78.971(8)	35 **Br** Bromine [79,901,79.907]	36 **Kr** Krypton 83.798(2)
37 **Rb** Rubidium 35.4678(3)	38 **Sr** Strontium 87.82(1)	39 **Y** Yttrium 38.90584(2)	40 **Zr** Sironium 91.224(2)	41 **Nb** Niobium 92.90637(2)	42 **Mo** Motybdenum 85.95(1)	43 **Tc** Technetium <98>	44 **Ru** Ruthenium 101.07(2)	45 **Rh** Rhodium 101.90550(2)	46 **Pd** Palladium 106.42(1)	47 **Ag** Silver 107.8882(2)	48 **Cd** Cadmium 112.414(4)	49 **In** Indium 114.818(1)	50 **Sn** Tin 118.710(7)	51 **Sb** Antimony 121.760(1)	52 **Te** Tellurium 127.60(3)	53 **I** Iodine 126.96447(3)	54 **Xe** Xenon 131.293(8)
55 **Cs** Cesium 132.90545196(6)	56 **Ba** Banum 137.327(7)	57-71	72 **Hf** Hafnium 178.49(2)	73 **Ta** Tantalum 180.94788(2)	74 **W** Tungsten 189.84(1)	75 **Re** Rhenium 186.207(1)	76 **Os** Osmium 190.23(3)	77 **Ir** Iridium 192.217(3)	78 **Pt** Platinum 195.084(9)	79 **Au** Gold 196.966569(5)	80 **Hg** Mercury 200.592(3)	81 **Ti** Thallium [204.382;204.385]	82 **Pb** Lead 207.2(1)	83 **Bi** Bismuth 208.98040(1)	84 **Po** Polonium <209>	86 **At** Astatine <210>	86 **Rn** Radon <222>
87 **Fr** Francium <223>	88 **Ra** Radium <226>	89-103	104 **Rf** Rutherfordium <267>	105 **Db** Dubnium <268>	106 **Sg** Seaborgium <271>	107 **Bh** Bohrium <272>	108 **Hs** Hassium <270>	109 **Mt** Meitnerium <278>	110 **Ds** Darmstadium <281>	111 **Rg** Roentgenium <280>	112 **Cn** Copernicium <285>	113 **Uut** Ununtrium unknown	114 **Fl** Flerovium <289>	115 **Uup** Ununpentium unknown	116 **Lv** Livermorium <293>	117 **Uus** Ununsentium unknown	118 **Uuo** Ununoctium unknown

Lanthanide Series	57 **La** Lanthanum 138.90547(7)	58 **Ce** Cerium 140.116(1)	59 **Pr** Praseodymium 140.90768(2)	60 **Nd** Neodymium 144.242(3)	61 **Pm** Promethium <145>	62 **Sm** Samarium 150.38(2)	63 **Eu** Europium 151.984(1)	64 **Gd** Gadolinium 157.25(3)	65 **Tb** Terbium 158.92535(2)	66 **Dy** Dyspresium 162.500(1)	67 **Ho** Holmium 164.83033(2)	68 **Er** Erbium 167.258(3)	69 **Tm** Thulium 168.93422(2)	70 **Yb** Ytterbium	71 **Lu** Lutetium 174.9658(1)
Actinide Series	89 **Ac** Actinium <227>	90 **Th** Thorium 232.0377(4)	91 **Pa** Protactinium 231.03588(2)	92 **U** Uranium 238.02891(3)	93 **Np** Neptunium <237>	94 **Pu** Plutonium <244>	95 **Am** Americium <243>	96 **Cm** Curium <247>	97 **Bk** Berkelium <247>	98 **Cf** Californium <251>	99 **Es** Einsteinium <252>	100 **Fm** Fermium <257>	101 **Md** Mendelevium <258>	102 **No** Nobelium <258>	103 **Lr** Lawrencium <262>

Characteristics of Modern Periodic Table : Modern Periodic Table is divided into groups and periods. Main characteristics of Modern Periodic Table are :

Groups

(i) Vertical columns in Modern Periodic Table are known as groups or families.

(ii) There are 18 groups in Modern Periodic Table and they are numbered from 1 to 18.

(iii) Each group consists of a number of elements having same valence electrons in its outermost shell.

(iv) In each group elements are classified into four main blocks, that are *s*, *p*, *d* and *f* blocks. This classification is based on the electronic configuration of the atoms and the name of orbital which receives the last electron.

(v) Group 1, 2, 13, 14, 15, 16, 17 and 18 are known as normal or representative elements, while groups 1 and 2 in Modern Periodic Table are known as ***s*-block elements**.

(vi) Group 1 elements are called **alkali metals** and group 2 elements are known as **alkaline earth metals.**

(vii) Group 3 to 12 are known as **transition elements** and their last and second last shells are incomplete.

(viii) Group 16 elements are called **chalcogens** and group 16 elements are called **halogens**

(ix) Group 18 elements are called noble gases, and their valence shell is completely filled.

Periods

(i) Horizontal rows in Modern Periodic Table are known as periods.

(ii) There are seven periods in Modern Periodic Table and they are numbered from 1 to 7.

These period number corresponds to the highest principal quantum number of the elements in periodic table.

(iii) 14 elements of sixth period are called lanthanides and 14 elements of seventh period are called the actinides. They are placed separately at the bottom of the table.

(iv) First period consists of two elements, second and third period consist of eight elements each, fourth and fifth period consist of 18 elements, sixth period consists of 32 elements and seventh period is incomplete.

Period	Valence shell	Type of Period	Number of Elements	Atomic Number of Elements
1st Period	$n = 1$	Short period	2	Atomic number 1 and 2
2nd Period	$n = 2$	Short period	8	Atomic number 3 to 10
3rd Period	$n = 3$	Long period	8	Atomic number 11 to 18
4th Period	$n = 4$	Long period	18	Atomic number 19 to 36
5th Period	$n = 5$	Long period	18	Atomic number 37 to 54
6th Period	$n = 6$	Long period	32	Atomic number 55 to 86
7th Period	$n = 7$	Incomplete	23	Atomic number 87 to 109

Merits of Modern Periodic Table

Modern Periodic Table explains the limitations of Mendeleev's periodic table and has many advantages. Some of them are :

(1) Position of isotopes : As Modern Periodic Table is based on atomic number, isotopes can be placed at one place in same group in Modern Periodic Table. In this way, he overcame the drawbacks of Mendeleev's periodic table.

For example : Three isotopes of hydrogen *i.e.*, $_1H^1$, $_1H^2$ and $_1H^3$ can be placed at the same place as that of hydrogen in the Modern Periodic Table.

(2) Anomalous position of some pairs of elements : In Mendeleev's Periodic table, some of the elements having higher atomic masses were placed before the elements having lower atomic mass. This efect of Mendeleev's periodic table was overcome in Modern Periodic table since it is based on atomic number.

Example : In Mendeleev periodic table, cobalt with atomic mass 58.93 u was placed before nickel having lower atomic mass 58.71 u but according to Modern Periodic table, atomic number of Co is 27 and that of Ni is 28. Thus, Ni should be placed after Co.

(3) Prediction of new elements : In Mendeleev's periodic table, it is not possible to predict how many new elements could be discovered between two known elements on the basis of atomic mass but in Modern Periodic table, the position and properties of new elements can be predicted easily on the basis of their atomic numbers or electronic configuration.

(4) Easy to remember and reproduce : The elements in Modern Periodic table are arranged in increasing order of their atomic numbers. Since atomic numbers increases in serial order, therefore, it is much easy to remember and reproduce Modern Periodic table.

Demerits of Modern Periodic Table

In spite of so many advantages of Modern Periodic Table, it has one disadvantage that the position of hydrogen is not certain as hydrogen shows similarity both with group 1(Alkali metals) and group 17 (Halogens) elements. At the same time, hydrogen differs from both the alkali metals and halogens in the other properties that oxides of H *i.e.*, water is neutral, oxides of alkali metals are basic while that of halogens are acidic. **Thus, the position of hydrogen is still controversial in Modern Periodic Table.**

5.5. SPECIAL REFERENCE TO ALKALI METALS (GROUP 1), ALKALINE EARTH METALS (GROUP 2), HALOGEN (GROUP 17) AND ZERO GROUP (GROUP 18)

The main group elements of the Periodic table are groups 1, 2 and 13 to 18. Elements in these groups are collectively known as main group or representative elements. These groups contain the most naturally abundant elements, comprising 80 percent of the earth's crust and are important for life.

Alkali Metals : (Group 1)

The Alkali metals comprises group 1 of the periodic table and they consists of lithium (Li), sodium (Na), rubidium (Rb), caesium (Cs), and francium (Fr). This group lies in the *s*-block of the periodic table as all the alkali metals have their outermost electron in *s*-orbital.

Properties

(1) The alkali metals are silver-coloured (caesium has a golden tinge), soft, low density metals. They can be cut very easily with butter knife and caesium can melt in the palm.

(2) All the alkali metal elements have one valence electron in its outermost shell which is very easily lost to form an ion with a positive charge.

(3) They have the lowest ionization energies in their respective periods. Ionization energy decreases down the group that is lithium has least ionization energy while, caesium has most ionization energy.

(4) These are the most active metals. Due to their activity they occur naturally in ionic compounds but not in their elemental state.

(5) These elements have largest atomic sizes in respective periods. Atomic size increases down the group.

Alkaline Earth Metals (Group 2)

The Alkaline earth metals comprises group 2 A of the periodic table and they consists of beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra). Alkaline earth metals have only two electrons in their outermost shell.

Properties

(1) The alkaline earth metals are silvery coloured, soft, low-density metals.

(2) Unlike the alkali metals, the earth metals have a smaller atomic size. However, atomic size increases down the group.

(3) These metals may also form ionic and other compounds and have a charge of +2. Beryllium is the least metallic element in the group and tends to form covalent bonds in its molecules.

(4) These are less reactive than alkali metals.

(5) Ionization energy decreases down the group.

(6) Alkaline earth metals have low melting point and boiling point.

Halogens (Group 17)

Halogens comprises group 17 of the periodic table and they consists of the five nonmetal elements fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (As). The term 'halogen' means 'salt-former' and compounds that contain one of the halogens are salts. They have seven electrons in their outermost shell, thus it is easy to gain one electron and convert into ion with –1 charge.

Properties

(1) The physical properties of halogens vary significantly as they exist as solids, liquids, and gases at room temperature. Fluorine and chlorine exist as gases while, bromine as a liquid, and iodine as a solid at room temperature.

(2) In general, halogens are very reactive, especially with the alkali metals and alkaline earth metals of group 1 and 2 respectively, with which they form ionic compounds.

(3) Fluorine is the most reactive and iodine is least reactive. The reactivity decreases as we go down in the group.

(4) The ionization energy of halogens is very high. Also ionization energy decreases down the group.

(5) Halogens have most negative electron gain enthalpy.

Nobel Gases (Group 18)

Nobel gases comprise the group 18 of the periodic table and they consists of helium (He), Neon (Ne), argon (Ar), krypton (Kr), xenon (Xe) and radon (Rn). Noble gases have completely filled outermost shell.

Properties

(1) The noble gases are all non-metals and are characterized by having completely filled shells of electrons.

(2) Because noble `gases' outer shells are full, they are extremely stable, tending not to form chemical bonds and having a small tendency to gain or lose electrons.

(3) The noble gases have very low boiling and melting points and are all gases at room temperature.

(4) The noble gases have little tendency to lose or gain electrons. Noble gases have zero electron affinity.

INTEXT QUESTIONS

1. Nobel gases belongs to which group ?
2. An element with atomic number 30 belongs to which of the group ?
3. How many elements are present in the 3rd period of the Modern Periodic Table ?
4. How do we define halogens ? Give examples.
5. How many groups and periods are there in Modern Periodic Table ?

SUMMARY

- Classification of elements was done to make their study easier.
- **Dobereiner's Law of Triad :** According to the law, the atomic mass of the middle element of a triad was nearly equal to the arithmetic mean of atomic masses of other two elements.
- **Newland's Law of Octaves :** "If elements were arranged in horizontal rows in order of ascending order of their relative atomic masses, then the properties of eighth element were repetition of the properties of first element". He compared this repetition with the eighth note on a musical scale.
- **Mendeleev** arranged 63 elements known at that time in the periodic table. According to Mendeleev "the properties of the elements are a periodic function of their atomic masses." The table consists of eight vertical column called "**Groups**" and horizontal rows called "**Periods**".
- The horizontal rows present in the periodic table are called periods, there are seven periods in the periodic table.
- The vertical columns present in it are called groups. These are nine in number and are numbered from I to VIII and Zero.
- Moseley gave Modern Periodic Law which states that "The physical and chemical properties of the elements are periodic function of their atomic number".

- Vertical columns in Modern Periodic Table are known as groups or families. There are 18 groups in Modern Periodic Table.
- Horizontal rows in Modern Periodic Table are known as periods. There are seven periods in Modern Periodic Table and are numbered from 1 to 7.
- Group 1 elements are called alkali metals, group 2 elements are known as alkaline earth metals, Group 17 elements are called halogens, while Group 18 elements are called noble gases. Noble gases have completely filled valence shell.

EXERCISE

A. VERY SHORT ANSWER TYPE QUESTIONS

1. State Newland's Law of Octave.
2. How many elements were known at the time of Newlands ?
3. A, B and C are the elements of a Dobereiner's triad. If the atomic mass of A is 8 and that of C is 40, what should be the atomic mass of Y ?
4. State the Modern Periodic Law. Who gave Modern Periodic Law ?
5. How many periods and groups are there in the long form of Periodic Table ?
6. How many elements are known to us at present ?
7. What was the main drawback of Dobereiner's Triad ?
8. Name the first element of Newland's octaves.
9. What is the fundamental property of elements according to Mendeleev ?
10. Chlorine, bromine and iodine forms a Dobereiner triad. The atomic masses of chlorine and iodine are 35 and 127. What will be the atomic mass of bromine ?
11. State the number of elements in period 1, period 2 and period 3 of the periodic table.
12. Name the elements present in group 1.
13. Name the elements present in longest period of Periodic table.
14. Predict the group of an element X if its atomic number is 16.
15. Which is the longest period in the long form of periodic table ?

B. SHORT ANSWER TYPE QUESTIONS

1. Why are noble gases discovered late ?
2. What are the reasons that Mendeleev left some gaps in Periodic Table ?
3. What are the limitations of Newland's Law of Octaves ?
4. Li, Be, B, C, O, F, Ne : These elements belongs to short period of the Periodic Table. Based on the given information answer following questions.
   (i) To which period do these elements belongs to? Which one is the missing element and where should it be placed ?
   (ii) Which one of the above elements belongs to the halogen group ?
5. Cl, Li, Ca, Br, Na, Sr, I, K and Ba. From the given elements :
   (i) Make three triads in total
   (ii) Name the family of each triad
6. The atomic number of an element A is 5. Identify the group and the period to which it belongs.
7. Why do the elements placed in the same group of the periodic table have similar chemical properties ?
8. Which one is odd in the given elements and why? Ca, Mg, Na and Be.
9. Give reason :
   (a) Noble gases have high Ionization potential.
   (b) Noble gases have zero electron affinity.
10. State the limitations of Dobereiner's Triad ?
11. What are the drawbacks of Mendeleev's Periodic table ?
12. (a) What do we call the vertical columns in a Periodic table ?
    (b) State some of the properties of alkaline earth metals ?
    (c) State some of the properties of noble gases.

C. LONG ANSWER TYPE QUESTIONS

1. The following questions refer to the periodic table :
 (i) Name the second last element of the period 3.
 (ii) How many elements are in the second period?
 (iii) Name the element which has the highest electron affinity.
 (iv) Name the element which has the highest electro-negativity.
 (v) Name the element which may be placed in group 1 but is not a metal.

2. The elements of the second group of the periodic table are beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra).
 (a) These elements belongs to which group?
 (b) In the given elements, which element has highest atomic size ?
 (c) Which among these elements is the least metallic ?
 (d) Which among these element has highest ionisation energy ?
 (e) What is the nature of the elements : metallic or non-metallic ? Why ?

3. What are the merits of long form of Periodic Table ?

4. (a) Na, Si, Cl : are these a Dobereiner triad? Explain.
 (Na-23, Si-28, Cl-35)
 (b) Explain two criteria's which were used by Mendeleev to classify elements ?
 (c) Carbon has two isotopes, where are these placed in the Modern Periodic table ?

D. FILL IN THE BLANKS

1. There are ……….. elements in Mendeleev periodic table.
2. Three elements arranged in increasing order of atomic masses by Dobereiner are called a ………
3. According to Law of Octaves, the elements similar to Hydrogen is ……..
4. Group 1 elements are known as…….. metals.
5. Electron affinity of noble gases is ……..
6. Modern periodic table is based on atomic ……..
7. Vertical columns in a Periodic Table are known as…….

[**Ans :** (1) 63, (2) Triad, (3) Fluorine, (4) Alkali, (5) Zero, (6) number, (7) Groups.]

E. TRUE OR FALSE

1. Ionisation energy of alkali metals is very low.
2. Alkaline earth metals shows a valency of 2.
3. Modern periodic law is given by Thomson.
4. Li, Na and I forms Dobereneir triad
5. Mendeleev's period law was based on atomic mass.
6. Very short period contains two elements only.
7. Atomic radius of Na is more than K.

[**Ans :** (1) True, (2) True, (3) False, (4) False, (5) True, (6) True, (7) False.]

F. MATCH THE COLUMN

Column A	Column B
1. Group 1	(a) Alkaline earth metals
2. Group 2	(b) Noble gases
3. Group 17	(c) Alkali metals
4. Group 18	(d) Chalcogens
5. Group 16	(e) Halogen

[**Ans :** 1. (c), 2. (a), 3. (e), 4. (b), 5. (d).]

G. MULTIPLE CHOICE QUESTIONS

1. Which of the following is a Dobereiner's Triad ?
 (a) Na, Cl, Ca (b) Li, Br, Sr
 (c) Cl, Br, I (d) K, I, Ba

2. If the atomic mass of first and third element of Dobereiner's Triad is 40 and 137, what will be the expected mass of middle element ?
 (a) 80.9 (b) 88.5
 (c) 90.1 (d) 125.5

3. According to Mendeleev's Periodic Law, the physical and chemical properties of elements are periodic function of their
 (a) atomic mass (b) atomic number
 (c) atomic size (d) valency

4. How many groups and periods are there in modern periodic table ?

(a) 16 and 7

(b) 6 and 17

(c) 18 and 7

(d) None of the above

5. Group 18 elements are known as :

(a) Halogens

(b) Alkali metals

(c) Alkaline earth metals

(d) Noble gases

6. Which group in periodic table has maximum number of elements ?

(a) Group 1 (b) Group 2

(c) Group 3 (d) Group 4

7. Law of Octaves was stated by :

(a) Dobereiner (b) Newland

(c) Mendeleev (d) Moseley

8. Which of the following is a set of alkaline earth metals ?

(a) Ca, Sr, Ba (b) Na, Li, K

(c) Na, K, Ca (d) Na, K, Rb

9. Long form of Periodic table was given by :

(a) Moseley

(b) J. J. Thomson

(c) Rutherford

(d) None of the above

10. Which of the following set of elements belongs to same group?

(a) 9 and 11 (b) 11 and 12

(c) 9 and 17 (d) 12 and 13

11. Which of the following shows correct pair according to Newland's Law of Octaves?

(a) H and O (b) O and S

(c) F and S (d) Cl and Ca

12. Which of the following pair of elements belongs to same group?

(a) Cu, Ag, Au

(b) Li, Be, Al

(c) Na, Mg, Al

(d) O, S, Cl

[**Ans :** 1. (c), 2. (b), 3. (a), 4. (c), 5. (d), 6. (a), 7. (b), 8. (a), 9. (d), 10. (c), 11. (b), 12. (a).

✦ ✦ ✦

CHAPTER 6

STUDY OF THE FIRST ELEMENT : HYDROGEN

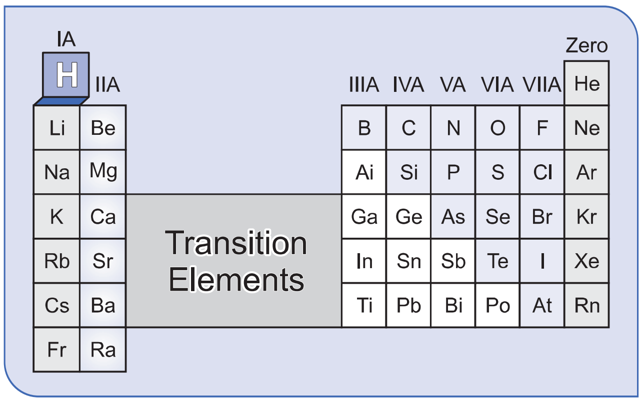
LEARNING OUTCOMES

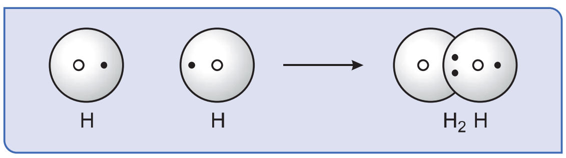
6.1. Position of Hydrogen in the Periodic Table
6.2. Similarities between Hydrogen and Alkali metals
6.3. Similarities between Hydrogen and Halogens
6.4. Discovery of Hydrogen
6.5. Occurence
6.6. Preparation of Hydrogen
6.7. Activity Series
6.8. Displacement of Hydrogen by Alkalis
6.9. Laboratory Preparation of Hydrogen
6.10. Manufacture of Hydrogen : Industrial methods
6.11. Properties of Hydrogen
6.12. Uses of Hydrogen
6.13. Oxidation and reduction

6.1. POSITION OF HYDROGEN IN THE PERIODIC TABLE

Symbol – H **Molecular formula** – H_2
Valency – 1 **Relative atomic mass** = 1
Relative molecular mass = 2
Electronic configuration = 1 (K shell)
Nature–Neutral gas **Solubility**-sparingly soluble in water.

The elements in the periodic table are arranged as per their atomic numbers, Hydrogen is the first element in the periodic table placed in 1^{st} Period and IA Group. Its atomic number is 1 containing only one electron in its outermost shell. Hydrogen has the simplest atomic structure amongst all the elements. It exists as a diatomic molecule and is called as Dihydrogen. Hydrogen exhibits similarities with Alkali metals (Group IA) and Halogens (Group VIIA). Alkali metals lose one electron to attain stability while Halogens gain one electron to attain stability. Hydrogen can lose an electron to form a positive ion (H^+- Hydronium) or gain an electron to form a negative ion (H^--Hydride). However, the position of Hydrogen in the periodic table has been controversial.





6.2. SIMILARITIES BETWEEN HYDROGEN AND ALKALI METALS

(a) Electronic configuration : Alkali metals like hydrogen have only one electron in their valance shell.

Group I : The alkali Metals

		Electronic configuration	
Li	Lithium	2, 1	Li - 3
Na	Sodium	2, 8, 1	Na - 11
K	Potassium	2, 8, 8, 1	K - 19
Rb	Rubidium	2, 8, 18, 8, 1	Rb - 37
Cs	Cesium	2, 8, 18, 18, 8, 1	Cs - 55
Fr	Francium	2, 8, 18, 32, 18, 8, 1	Fr - 87

(b) Valency : All alkali metals have valency 1.

(c) Ion formation : Each of the alkali metals and Hydrogen forms a cation by losing one electron. Since alkali metals lose an electron to attain stability, they are electropositive in nature.

$$H \rightarrow H^{+} + e^{-}$$
$$Li \rightarrow Li^{+} + e^{-}$$
$$Na \rightarrow Na^{+} + e^{-}$$

(d) Reducing nature : Both hydrogen and alkali metals are good reducing agents while they themselves get oxidised. Consider below given examples :

$$CuO + H_2 \rightarrow Cu + H_2O$$
$$CuO + 2Na \rightarrow Cu + Na_2O$$

(e) Burning : Hydrogen and alkali metals burns in the presence of oxygen to form various oxides.

For example : Hydrogen burns in oxygen to form water.

$$2H_2 + O_2 \rightarrow 2H_2O$$

Lithium burns in oxygen to form a monoxide

$$4Li + O_2 \rightarrow 2Li_2O$$

Sodium burns in oxygen to form a peroxide

$$2Na + O_2 \rightarrow 2Na_2O_2$$

Potassium, Rubidium and Caesium forms superoxides having the general formula MO_2, where M stands for metal.

Alkali metals are always stored in non-reactive organic solvents such as kerosene to prevent them from coming in contact with air because they rapidly react with oxygen, moisture and carbon dioxide present in air to form oxides, hydroxides and carbonates respectively.

Example :

$$4Na + O_2 \rightarrow 2Na_2O$$
$$Na_2O + H_2O \rightarrow 2NaOH$$
$$2NaOH + CO_2 \rightarrow Na_2CO_3 + H_2O$$

(f) Combination with non-metals : Hydrogen and alkali metals reacts with non-metals like sulphur, chlorine and oxygen to give various compounds like H_2O, Na_2O, H_2S, Na_2S, HCl, NaCl, etc.

6.3 SIMILARITIES BETWEEN HYDROGEN AND HALOGENS

Atomic Number	Name	Electronic Configuration				
9	Fluorine	2	7			
17	Chlorine	2	8	7		
35	Bromine	2	8	18	7	
53	Iodine	2	8	18	18	7

(a) Electronic Configuration : Halogens have 7 electrons in their outermost shell, so they lose one electron to attain the electronic configuration of the nearest noble gas in their valence shell. Similarly hydrogen can gain stability by accepting one electron.

(b) Ion formation : Each of the halogens and Hydrogen atoms forms an anion by gaining one electron. Since halogens accept an electron to attain stability, they are electronegative in nature.

$$H + e^{-} \rightarrow H^{-}$$
$$F + e^{-} \rightarrow F^{-}$$
$$Cl + e^{-} \rightarrow Cl^{-}$$

(c) Physical state : Like halogens, Hydrogen is also a gas.

(d) Atomicity : Like halogens, Hydrogen also exists in the form of diatomic molecule. For example : H_2, Cl_2 , F_2 etc.

However, due to all these similarities hydrogen differs from Alkali metals and Halogens with respect to the following properties :

(i) Hydrogen has only one shell but alkali metals and halogens have multiple shells.

(ii) Oxides of hydrogen are neutral in nature, but oxides of halogens are acidic in nature and oxides of alkali metals are basic in nature.

(e) Valency : Both Hydrogen and Halogens have valency 1.

(f) Non-metallic character : Like Halogen, hydrogen is non-metallic in Nature.

6.4. DISCOVERY OF HYDROGEN

In 1671, **Robert Boyle** discovered and proposed the reaction between iron filings and dilute acids.

As a result of this reaction hydrogen gas was produced. In 1766, **Henry Cavendish** recognized Hydrogen gas as a discrete substance, by naming it as 'inflammable air'. Further in 1781, he found that the gas produces water when burned. Henry Cavendish is usually given credit for the discovery of Hydrogen as an element. In 1783, **Antoine Lavoisier** gave this element the name hydrogen when he and Laplace reproduced Cavendish's finding that water is produced when hydrogen is burned.

6.5. OCCURRENCE

Hydrogen occurs in free state as well as in a combined state. The sun and other stars are composed of mainly hydrogen. Hydrogen is mainly composed of three isotopes protium, deuterium, and tritium. Earth's crust contains about 0.98% and atmosphere contains about 0.01% of hydrogen gas. Volcanic gases are known to contain 0.025% and atmosphere of sun contains 1.1% of hydrogen.

(a) Plant and Animal cells are made up of hydrogen with carbon, oxygen and other elements.

(b) Hydrogen is a key component in acids, alkalis, water, hydrocarbons, organic compounds and nutrients.

6.6. PREPARATION OF HYDROGEN

Hydrogen from water : Water contains about 11.10% of hydrogen by volume. This hydrogen can be liberated when water reacts with metals or by electrolysis.

(a) Reaction of metals with cold water : All the alkali metals react vigorously with cold water. Alkali metals react with cold water to form oxides and hydroxides with the liberation of hydrogen gas. Such reactions are exothermic reactions.

(i) Potassium : When potassium is added to water, it melts and floats on the surface of water forming a silvery grey globule. This reaction is highly exothermic and vigorous. Bubbles of Hydrogen gas are seen and a colourless solution is formed. During the reaction, the formed hydrogen ignites instantly. The metal catches fire, with sparks and a lilac flame. Red litmus paper turns blue if it is introduced in the solution.

Reaction : $2K + H_2O \rightarrow 2KOH + H_2$

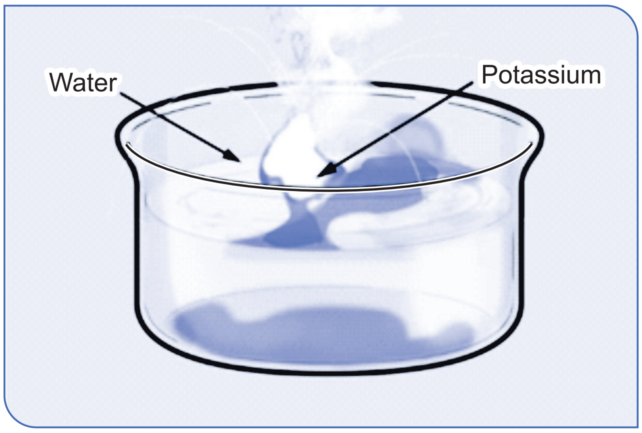


Fig. 6.1 : Reaction between potassium and cold water

(ii) Sodium : When sodium is added to water, it melts and forms a silvery globule that darts on the surface of water. This reaction is less exothermic and less vigorous as compared to that of potassium. It catches fire and burns with a golden yellow flame. The solution is alkaline and turns red litmus paper blue if it is introduced in the solution.

Reaction : $2Na + H_2O \rightarrow 2NaOH + H_2$

Fig. 6.2 : Reaction between sodium and cold water

(iii) Calcium : Calcium reacts slowly with water to give Calcium hydroxide, and Hydrogen gas. The calcium metal sinks in water and bubbles of hydrogen are stuck to the surface of Calcium. The reaction is less vigorous and exothermic than Sodium and Potassium. Red litmus paper turns blue if it is introduced in the solution.

Reaction : $Ca + H_2O \rightarrow Ca(OH)_2 + H_2$

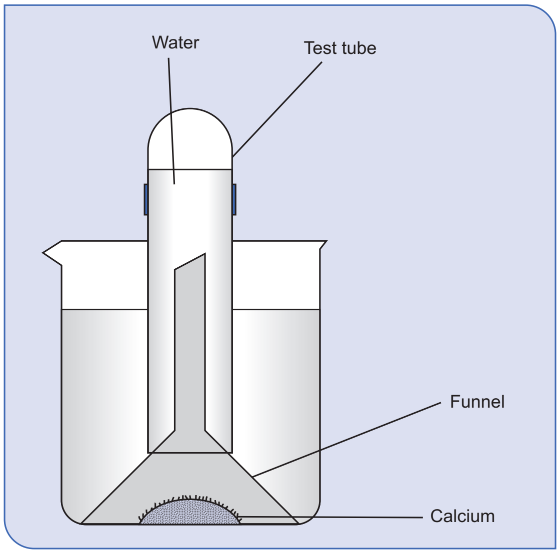


Fig. 6.3 : Reaction between calcium and cold water

(b) Reaction of metals with hot water or steam: Magnesium, Aluminium, Zinc and Iron does not react with cold water. Yet, they react with hot water or steam to form oxides and hydroxides and liberates hydrogen.

(i) Magnesium : It reacts slowly with hot water to give Hydrogen gas and Magnesium hydroxide, which is slightly soluble in water.

Reaction : $Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$

Hydrogen gas is evolved during the reaction. The resulting alkaline solution will turn red litmus blue because of the formation of magnesium hydroxide.

However, Magnesium reacts with steam to give magnesium oxide and liberates Hydrogen.

Reaction **:** $Mg + H_2O \rightarrow MgO + H_2$

(ii) Aluminium : Aluminium reacts readily with steam to give aluminium oxide and hydrogen

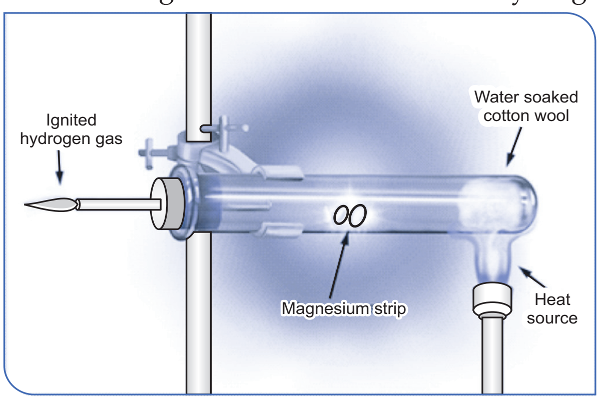


Fig. 6.4 : Reaction between Aluminium and Cold water

gas. Coating of aluminium oxide does not allow further reaction, thus it protects metal from other reactions. However, at high temperature this coating breaks and liberates more Hydrogen.

Reaction : $2Al + 3H_2O \rightarrow Al_2O_3 + 3H_2$

(iii) Zinc : Zinc reacts with steam to form zinc oxide and hydrogen gas. The reaction is less vigorous compared to that of Magnesium. Zinc oxide is yellow in colour when it is hot but turns white when it is cooled.

Reaction : $Zn + H_2O \rightarrow ZnO + H_2$

(iv) Iron : Red hot iron reacts with steam to give ferro-ferric oxide and hydrogen gas. It is a reversible reaction. A **reversible reaction** is a reaction where the reactants form products, which react together to give the reactants back. In this case, if the formed ferro-ferric oxide is not removed quickly, the oxide formed is reduced back to iron.

Reaction : $3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$

As the reaction proceeds reactants are consumed and simultaneously products are formed. As a result the rate of forward reaction increases and the rate of backward reaction decreases.

A stage in the reaction is seen where the rate of backward and forward reaction is same and the concentration of reactants and products are same. This stage is called as **equilibrium**. In this reaction, equilibrium is observed at 700°C.

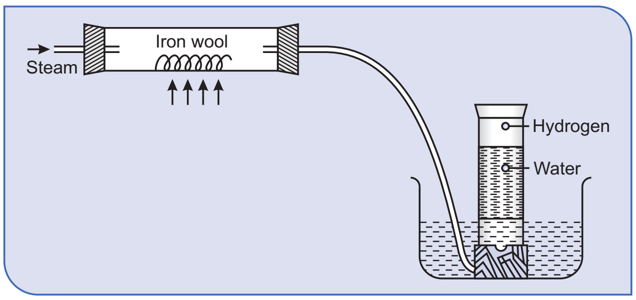


Fig. 6.5 : Reaction between Iron and Cold water

(c) Reaction of steam on non-metals :

Carbon : When steam is passed over red-hot coke, water gas is formed.

$$C + H_2O \rightarrow \underbrace{CO + H_2}_{\text{Water gas}}$$

INTEXT QUESTIONS

1. Justify the position of Hydrogen in the periodic table.
2. State any two differences between alkali metals and Hydrogen.
3. State any two similarities between Halogens and Hydrogen.
4. Hydrogen can form a positive or a negative ion. Justify.
5. Describe the occurrence of Hydrogen in its free state.
6. Hydrogen shows similarities with alkali metals and halogens. However, oxides of hydrogen, alkali metals and halogens differ. Explain this statement.

6.7. ACTIVITY SERIES

Reactivity series or **activity series** is a progression of a series of metals, arranged by their reactivity from highest to lowest. More reactive metal displaces less reactive metal.

Some of the salient features of activity series are :

(i) On moving down in the reactivity series the electropositive character of elements decreases.

(ii) In the same way, the reducing power of metal decreases down the series. Potassium is the strong reducing agent.

Reactivity Series of Metals

	Metal	Symbol	
These metals are more reactive than hydrogen	Potassium	K	(Most reactive metal)
	Sodium	Na	
	Calcium	Ca	
	Magnesium	Mg	
	Aluminium	Al	
	Zinc	Zn	
	Iron	Fe	
	Tin	Sn	
	Lead	Pb	
	[Hydrogen]	[H]	
These metals are less reactive than hydrogen	Copper	Cu	
	Mercury	Hg	
	Silver	Ag	
	Gold	Au	(Least reactive metal)

(iii) Metals have tendency to get oxidised (lose electrons) and this tendency decreases down the series. Potassium is the most readily oxidised metal.

(iv) Metals which are placed above hydrogen have tendency to displace hydrogen from water and dilute acids, but the metals which are placed below hydrogen do not have tendency to displace hydrogen.

$$2Na + H_2O \rightarrow 2NaOH + H_2$$

$$Cu + H_2O \rightarrow \text{No reaction}$$

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

$$Cu + HCl \rightarrow \text{No reaction}$$

(v) Metals which are placed in the higher positions in the reactivity series displaces a metal placed below in order to form salt solutions. Displacement is more rapid if the difference in their position is greater.

$$CuCO_3 + Zn \rightarrow ZnCO_3 + Cu$$

(vi) The oxides of metals like K, Na, Ca, Mg, Al cannot be reduced by common reducing agents such as H_2, CO or C, but they can be reduced by electrolysis.

(vii) Oxides and nitrates of less reactive metals decompose to give metals.

$$2HgO \rightarrow 2Hg + O_2$$

$$2AgNO_3 \rightarrow 2Ag + 2NO_2 + O_2$$

(viii) Metals below copper do not rust easily as they occur in its pure form.

(ix) Metals occupying top positions in the series easily combine with more electronegative elements to form chlorides, oxides, sulphides, etc.

Displacement of Hydrogen from Dilute acids

Some metals react with dilute acids to form salts by displacing hydrogen. Metals which are placed above Hydrogen in the activity series are capable of displacing Hydrogen which is shown in the table-1 given ahead.

Table 1 : Reaction of Metals with acids, Products formed and observations

Metal	Acid	Products formed	Observation
2K	2HCl H_2SO_4	$KCl + H_2$ $K_2SO_4 + H_2$	These reactions are highly explosive and practically not feasible.
2Na	2HCl H_2SO_4	$NaCl + H_2$ $Na_2SO_4 + H_2$	
Ca	2HCl H_2SO_4	$CaCl_2 + H_2$ $CaSO_4 + H_2$	
Mg	2HCl H_2SO_4	$MgCl_2 + H_2$ $MgSO_4 + H_2$	These metals react with dilute acid (HCl or H_2SO_4) to form a salt and hydrogen gas.
3Al	H_2SO_4	$Al_2(SO_4)_3 + H_2$	
Zn	H_2SO_4	$ZnSO_4 + H_2$	
Fe	2HCl	$FeCl_2 + H_2$	
Ni	H_2SO_4	$NiSO_4 + H_2$	Reaction of Tin with dilute acid is very slow. Lead reacts with HCl and H_2SO_4 and forms lead chloride and lead sulphate respectively. Thus, the reaction does not proceed further.
Sn	2HCl	$SnCl_2 + H_2$	
Pb			
Hydrogen			
Cu	These elements do not displace Hydrogen from the acid. Hence no reaction takes place.		
Ag			
Pt			
Au			

Chemical Reactions :

$$\underset{\text{(dil.)}}{Mg + 2HCl} \rightarrow MgCl_2 + H_2$$

$$\underset{\text{(dil.)}}{2Al + H_2SO_4} \rightarrow Al_2(SO_4)_3 + H_2$$

$$\underset{\text{(dil.)}}{Zn + H_2SO_4} \rightarrow ZnSO_4 + H_2$$

$$\underset{\text{(dil.)}}{Fe + 2HCl} \rightarrow FeCl_2 + H_2$$

Reasons for not using dilute nitric acid for the preparation of hydrogen gas

(i) This is because nitric acid is a strong oxidising agent as it decomposes to yield nascent oxygen, and it oxidises with hydrogen formed to water. Only 1% dilute and cold nitric acid reacts with magnesium and manganese to liberate hydrogen gas.

$$2HNO_3 \rightarrow 2NO_2 + H_2O + [O]$$

However, metals like magnesium and manganese reacts with very dilute nitric acid to produce hydrogen gas in very little amount.

(ii) Nitric acid is very strong oxidising agent which reacts with the produced hydrogen gas in the reaction and oxidises that hydrogen gas into water. Because of this reason nitric acid is not used for the preparation of hydrogen gas. Instead it evolves nitric oxides such as NO_2, NO etc.

Amongst all the metals, Zinc is the most feasible metal to prepare Hydrogen because of following reasons :

1. Sodium and Potassium reacts violently with acid.
2. Aluminium forms a tough layer of its oxide on the metal surface and hence further reaction does not occur.
3. Calcium and Magnesium are expensive metals.
4. Iron contain a lot of impurities.
5. Lead reacts with HCl and H_2SO_4 and forms lead chloride and lead sulphate respectively. Thus, the reaction does not proceeds further.

6.8. DISPLACEMENT OF HYDROGEN BY ALKALIS

Some metals react with concentrated alkalis to form salts and hydrogen gas which is shown in below table 2.

Table 2 : Reaction of Metals with Alkali

Metal	Alkali (conc.)	Products formed
Zn	2NaOH	$Na_2ZnO_2 + H_2$ (sodium zincate)
Zn	2KOH	$K_2ZnO_2 + H_2$ (potassium zincate)
Pb	2NaOH	$Na_2PbO_2 + H_2$ (sodium plumbate)
2Al	$2NaOH + 2H_2O$	$Na_2AlO_2 + H_2$ (sodium aluminate)

However, some metals are amphoteric in nature. Oxides and Hydroxides of Zinc, Lead and Aluminium reacts with acids as well as alkalis to form salts with the evolution of hydrogen gas.

6.9. LABORATORY PREPARATION OF HYDROGEN

Procedure

Take some pieces of Zinc granules in a flask fitted with air tight rubber cork with two holes containing a thistle funnel and a delivery tube. Pour dilute hydrochloric acid through thistle funnel.

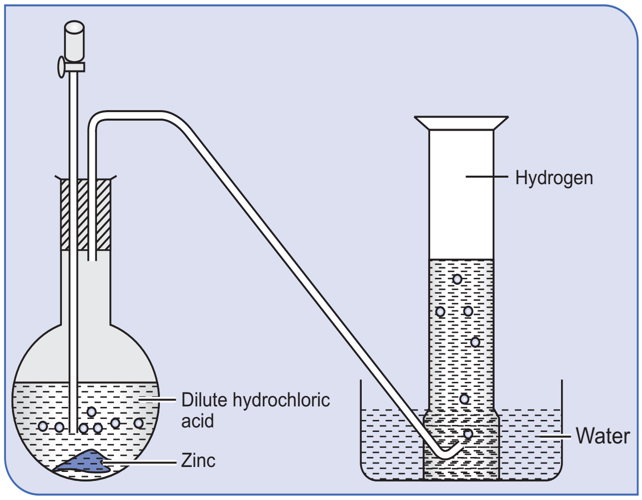


Fig. 6.6 : Laboratory preparation of hydrogen by the action of dilute acid on Zinc

Observation : Effervescence are formed and evolution of gas is observed. The hydrogen gas is lighter than air and forms an explosive mixture with air. Hydrogen is collected by the downward displacement of water because it is almost insoluble in water.

Chemical reaction :

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

Impurities present with Hydrogen

(i) Moisture
(ii) Arsine (AsH_3)
(iii) Phosphine (PH_3)
(iv) Carbon dioxide (CO_2)
(v) Sulphur dioxide (SO_2)
(vi) Hydrogen sulphide (H_2S)
(vii) Nitrogen dioxide (NO_2)

Purification

Hydrogen gas thus collected may contain impurities like H_2S, SO_2, PH_3, CO_2, water vapour, etc. These impurities are removed by passing hydrogen gas through three different reagents and a 'U' shaped tube containing anhydrous Calcium chloride. The diagram depicts the stepwise purification process of hydrogen gas. Following chemical reaction occur during the process and during these processes the impurities are also removed from hydrogen.

(i) Silver nitrate reacts with arsine to form an insoluble yellow precipitate of Silver arsenide, Silver nitrate and Nitric acid.

$$AsH_3 + 6AgNO_3 \rightarrow Ag_3As + 3AgNO_3 + 3HNO_3$$

Silver nitrate reacts with phosphine to form an insoluble yellow precipitate of Silver Phosphide Silver nitrate and Nitric acid.

$$PH_3 + 6AgNO_3 \rightarrow Ag_3P + 3AgNO_3 + 3HNO_3$$

Thus, Arsine and Phosphine are removed.

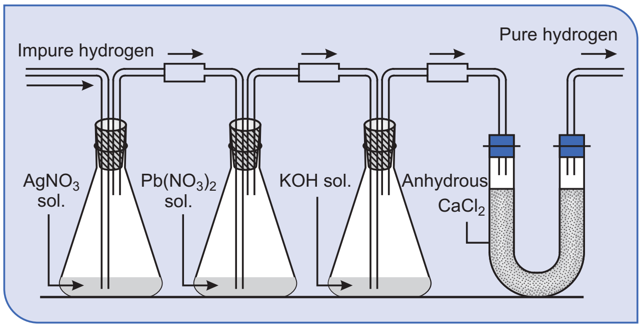


Fig. 6.7 : Purification of hydrogen gas

(ii) Lead nitrate reacts with Hydrogen sulphide to form an insoluble black precipitate of lead sulphide. Thus, Hydrogen sulphide is removed.

$$Pb(NO_3)_2 + H_2S \rightarrow PbS + 2HNO_3$$

(iii) Caustic potash solution removes SO_2,CO_2 and oxides of nitrogen.

$$SO_2 + KOH \rightarrow K_2SO_3 + H_2O$$
$$CO_2 + KOH \rightarrow K_2CO_3 + H_2O$$
$$2NO_2 + KOH \rightarrow KNO_2 + KNO_3 + H_2O$$

(iv) Anhydrous Calcium chloride absorbs moisture from the purified hydrogen, hence water vapour is removed.

Precautions

Following precautions should be taken while carrying out this process :

(i) The tip of the thistle funnel should be dipped below the level of the dilute acid in the round bottom flask else the hydrogen gas formed may escape through the round bottom flask.

(ii) Hydrogen gas should be collected after the apparatus is evacuated. All the air should be allowed to escape.

(iii) Flame source should not be kept near the apparatus setup as hydrogen gas is explosive.

(iv) Apparatus should be checked for any damages before starting the experiment to avoid gas leakage.

6.10. MANUFACTURE OF HYDROGEN : INDUSTRIAL METHODS

(a) Bosch Process : The Bosch process is used for the industrial manufacture of hydrogen gas.

Step 1 : Hydrogen is prepared by passing steam at 1200°C over a red hot coke. As result, a mixture is called water gas.

$$C + H_2O \xrightarrow{1270\ k} \underset{\text{(water gas)}}{CO + H_2}$$

Step 2 : In step 2 the produced water gas is mixed with excess of steam and passed through ferric oxide which acts as a catalyst and chromic oxide which acts as a promoter at about 450°C. The resultant reaction produces a mixture of carbon dioxide and hydrogen gas. This reaction is exothermic.

$$\underset{\text{(water gas)}}{CO + H_2} + H_2O \xrightarrow{\text{feeric chromate}} CO_2 + 2H_2 + \text{heat}$$

Step 3 : In this step carbon dioxide is separted by passing through cold water under pressure of 30 atm or caustic potash solution. The CO_2 dissolves in caustic potash solution leaving hydrogen.

$$2KOH + CO_2 \longrightarrow K_2CO_3 + H_2O$$

Finally this mixture is passed through a solution of ammoniacal cuprous chloride to remove unreacted carbon monoxide. Hence hydrogen gas is left behind.

$$CuCl + CO + 2H_2O \longrightarrow CuCl.\ CO.2H_2O$$

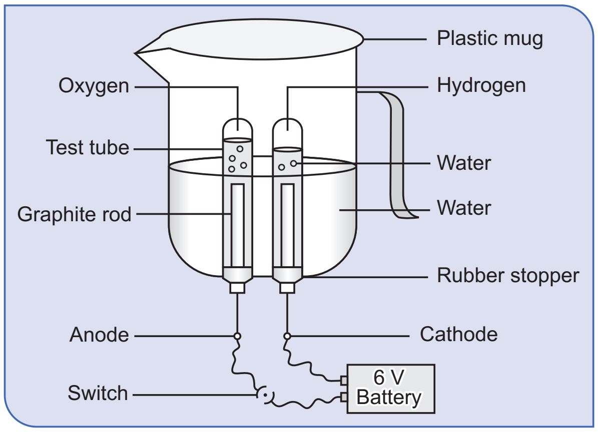


Fig. 6.8 : Preparation of hydrogen gas by electrolysis of water

(b) By electrolysis of water : Another method of producing hydrogen industrially is the electrolysis of water using platinum electrode. In this case water is acidified by adding few drops of sulphuric acid. When electric current is passed through the water it decomposes to hydrogen and oxygen. Hydrogen ions moves towards cathode by releasing electrons, while hydroxide ions move towards anode which gain released electrons from cathode.

Hydrogen gas is released at cathode and oxygen gas is relesed at anode. In the process of electrolysis oxygen gas and hydrogen gas are simultaneously collected and both the gases are pure. The level of oxygen gas is double than that of hydrogen gas.

$$2H_2O \xrightarrow{\text{Electricity}} 2H_2 + O_2$$

Mechanism :

$$H_2O \longleftrightarrow H^+ + OH^-$$

At cathode :

$$H^+ + e^- \longrightarrow H$$
$$H + H \longrightarrow H_2\uparrow$$

At anode :

$$OH^- - e^- \longrightarrow OH$$
$$OH + OH \longrightarrow H_2O + O$$
$$O + O \longrightarrow O_2\uparrow$$

(c) From natural gas : Hydrogen gas can also be produced by natural gas mainly known as methane.

Fossil fuels are the dominant source of industrial hydrogen. Hydrogen gas can be generated from natural gas with approximately 80% efficiency. For the process, steam reacts with methane temperature of about 700-1100°C. The resultant mixture is known as synthesis gas. The gas is than separate by the technique used in Bosch process. This is an endothermic reaction yielding Hydrogen.

$$CH_4 + H_2O\text{ (steam)} \xrightarrow{\text{Ni}} CO + 3H_2$$

(d) From Calcium hydride : Calcium hydride reacts with steam to give calcium hydroxide and hydrogen gas.

$$CaH_2 + H_2O \longrightarrow Ca(OH)_2 + 2H_2$$

6.11 PROPERTIES OF HYDROGEN

(a) Physical Properties :

(i) *Colour* : colourless

(ii) *Odour* : odourless

(iii) *Taste* : tasteless

(iv) *Toxicity* : non-toxic

(v) *Density* : 0.0899g/ml. it is 14.4 times lighter than air

(vi) *Solubility* : very slightly soluble in water

(vii) *Liquefaction* : liquefies below -240°C under pressure of about 20 atmospheres

(viii) *Occlusion/Adsorption* : Hydrogen is readily adsorbed by nickel, platinum or palladium.

(b) Chemical Properties : Hydrogen gas is highly flammable and burns in air at very high concentrations. The enthalpy of combustion for hydrogen is – 286 kJ/mol.

(i) Reaction of Hydrogen with non-metals :

(a) *Oxygen*

$$2H_2 + O_2 \longrightarrow 2H_2O + 286\text{ kJ/mol}$$

(b) *Chlorine* : Hydrogen gas can also be explode when mixed with chlorine in the presence of catalyst.

$$H_2 + Cl_2 \longrightarrow 2HCl$$

(c) *Nitrogen*

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

(d) *Sulphur*

$$H_2 + S \longrightarrow H_2S$$

(ii) Reaction of Hydrogen with metals (in presence of heat) :

(a) *Potassium*

$$2K + H_2 \longrightarrow 2KH$$

(b) *Sodium*

$$2Na + H_2 \longrightarrow 2NaH$$

(c) *Calcium*

$$Ca + H_2 \longrightarrow CaH_2$$

Calcium hydride further reacts with water to form calcium hydroxide with the evolution of hydrogen gas.

$$CaH_2 + 2H_2O \longrightarrow Ca(OH)_2 + 2H_2$$

(iii) Reaction of Hydrogen with metallic oxides :

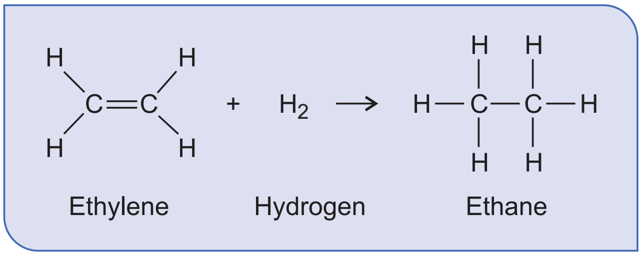
Hydrogen gas acts as a strong reducing reagent. It removes oxygen from the strongly heated metal oxide and itself gets oxidised to water.

For example :

$$Fe_2O_3 + 3H_2 \xrightarrow{\Delta} 2Fe + 3H_2O$$
$$Fe_3O_4 + 4H_2 \xrightarrow{\Delta} 3Fe + 4H_2O$$
$$PbO + H_2 \xrightarrow{\Delta} Pb + H_2O$$
$$CuO + H_2 \xrightarrow{\Delta} Cu + H_2O$$

(iv) Reaction with Hydrocarbons :

Hydrogenation is a process where hydrogen directly combines with unsaturated organic compounds in the presence of catalyst like nickel to form alkane.



6.12. USES OF HYDROGEN

- Hydrogen gas in the form of an oxy-acetylene flame is used for welding and cutting of metals. This flame can be used for melting platinum, quartz and for fusing alumina to produce synthetic rubies and sapphires.
- Hydrogen gas is used in the production of fuel gases like Coal gas and Water gas as it has high heat of combustion.
- It is also used to produce fuel oil, gasoline and ammonia.
- Liquid hydrogen can be used as a fuel for motor vehicles, fuels in rockets fuel cells which is a non-polluting gas.
- It is used to convert vegetable oil into vanaspati ghee, popularly known as 'Dalda' through the process of Hydrogenation.
- It is used in meteorological balloons to study weather conditions.
- Hydrogen gas is used for producing artificial petrol from coal. Passage of hydrogen under high pressure over powdered coal in the presence of a catalyst is known as hydrogenation of coal.
- Hydrogen gas is a good reducing agent thus it is used in extraction of metals. Hydrogen gas reduces heated metallic oxide to metals.

$$Fe_2O_3 + 3H_2 \xrightarrow{\Delta} 2Fe + 3H_2O$$

$$CuO + H_2 \xrightarrow{\Delta} Cu + H_2O$$

6.13. OXIDATION AND REDUCTION

(a) Oxidation :

Oxidation reaction involves the addition of oxygen or the removal of hydrogen from the substance.

$$2Mg + O_2 \longrightarrow 2MgO$$

(Mg is oxidised to MgO due to addition of oxygen)

$$H_2S + Cl_2 \longrightarrow 2HCl + S$$

(H_2S is oxidised to S due to removal of hydrogen)

Also addition of electronegative ion and removal of electropositive ion is seen during oxidation.

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

(Zinc is oxidised to zinc chloride due to the addition of electronegative chloride ion)

$$2KI + H_2O_2 \longrightarrow 2KOH + I_2$$

(potassium iodide is oxidised to iodine due to the removal of electropositive potassium)

Oxidation also involves loss of electron(s) from an atom or ion.

$$Zn \longrightarrow Zn^{2+} + 2e^-$$

Oxidising agent is a substance that has the ability to oxidise other substance by donating electrons, providing oxygen.

Test for oxidising agents

(i) When a strip of moist starch – iodide paper is dipped into a substance suspected to be an oxidizing agent, and blue-black colour appears, then the substance is confirmed to be an oxidising agent.

(ii) When hydrogen sulphide (H_2S) is reacted with a substance, and some yellow deposits (sulphur) are settled at the bottom of the container, then the substance is an oxidizing agent. H_2S is oxidized to sulphur.

(iii) On warming with concentrated hydrochloric acid (HCl) they liberate chlorine gas which finally bleaches moist red litmus paper.

(iv) On heating strongly, it liberates oxygen which rekindles the glowing splinter.

(b) Reduction :

Reduction reaction involves the removal of oxygen or the addition of hydrogen from a substance.

$$CuO + H_2 \longrightarrow Cu + H_2O$$

(CuO is reduced to Cu due to removal of Oxygen)

$$H_2S + Cl_2 \longrightarrow 2HCl + S$$

(Chlorine is reduced to HCl due to addition of Hydrogen)

Also addition of electronegative ion and removal of electropositive ion is seen during oxidation.

$$2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$$

(Mercuric chloride is reduced to mercurous chloride due to addition of electropositive Hg^+ ion)

$$2FeCl_3 + H_2S \longrightarrow 2FeCl_2 + 2HCl + S$$

(Ferric chloride is reduced to ferrous chloride due to removal of electronegative Cl^- ion)

Reduction also involves gain of electron(s) by an atom or ion.

$$Cu^{2+} + 2e^- \longrightarrow Cu$$

Reducing agent is a substance which has the ability to reduce other substance by accepting electrons, providing hydrogen.

Test for reducing agents

(i) They liberate brown fumes of nitrogen dioxide when heated with nitric acid.

(ii) They decolourise purple pink colour of dilute Potassium permanganate.

(iii) They change orange colour of acidified potassium dichromate to green.

(iv) They change light yellow colour of iron(III) salts to light green colour due to the formation of iron (II) salts.

(c) Redox Reaction : In some chemical reactions, oxidation and reduction occur together, means one substance gets oxidised while other gets reduced. Such reactions are called redox reactions. In the name 'redox', the term 'red' stands for reduction and 'ox' stands for oxidation. The common examples of redox reactions are : corrosion, respiration, working of battery cell.

1. $CuO + Mg \longrightarrow MgO + Cu$

$Cu^{2+} + Mg \longrightarrow Cu + Mg^{2+}$

$CuO \longrightarrow Cu$

(Copper oxide is reduced to copper)

$Cu^{2+} + 2e^- \longrightarrow Cu$ (reduction)

$Mg \longrightarrow MgO$

(Mg is oxidised to MgO)

$Mg \longrightarrow Mg^{2+} + 2e^-$ (oxidation)

2. $Zn + CuSO_4 \longrightarrow Cu + ZnSO_4$

$Zn \longrightarrow ZnSO_4$

(Zn is oxidised to $ZnSO_4$)

$Zn \longrightarrow Zn^{2+} + 2e^-$ (oxidation)

$CuSO_4 \longrightarrow Cu$

($CuSO_4$ is reduced to Cu)

$Cu^{2+} + 2e^- \longrightarrow Cu$ (reduction)

Some more examples of redox reactions are given below :

(i) $ZnO + C \longrightarrow Zn + CO$

(ii) $MnO_2 + HCl \longrightarrow MnCl_2 + H_2O + Cl_2$

(iii) $2\,FeCl_3 + H_2 \longrightarrow 2\,FeCl_2 + 2\,HCl$

(iv) $3Fe_3O_4 + 8\,Al \longrightarrow 9\,Fe + 4Al_2O_3$

Table 3 : Difference between Oxidation and Reduction.

	Oxidation (with respect to A)	Reduction (with respect to A)
(i)	Oxidation is the loss of electrons or hydrogen atoms or gain of oxygen atoms.	Reduction is the gain of electrons or hydrogen atoms or loss of oxygen atoms.
(ii)	Removal or loss of electrons $A \rightarrow A^+ + e^-$	Addition or gain of electrons $A + e^- \rightarrow A^-$
(iii)	Removal of Hydrogen $AH + B \rightarrow A + BH$	Addition of Hydrogen $A + BH \rightarrow AH + B$
(iv)	Addition of oxygen $A + BO \rightarrow AO + B$	Removal of oxygen $AO + B \rightarrow A + BO$
(v)	Energy gets released during oxidation	During reduction energy is gained.

where A, B = Metals, H = Hydrogen and O = Oxygen.

Examples of oxidation reduction reactions are corrosion, respiration, the operation of battery etc.

SUMMARY

- Hydrogen is the first element in the periodic table that is placed in 1^{st} Period and IA Group. Its atomic number is 1 containing only one electron.
- Hydrogen exhibits similarities and few differences with alkali metals (Group IA) and halogens (Group VIIA).
- Hydrogen can lose an electron to form a positive ion (H^+- Hydronium) or gain an electron to form a negative ion (H^-- Hydride).
- Activity series of metals is an arrangement of metals in the decreasing order of reactivity with water in the form of a series.
- Reactive metals react with cold water to form oxides and hydroxides and liberates hydrogen. Such reactions are exothermic reactions.
- Some metals react with dilute acids to form salts by displacing hydrogen. Metals which are placed above hydrogen in the activity series are capable of displacing hydrogen.
- Amphoteric metals react with acids as well as alkalis to form salts and liberates hydrogen.
- Hydrogen can be prepared by action of dilute acid on zinc, Bosch process or electrolysis of water and reaction with natural gas.

- Hydrogen is a reducing agent.
- Hydrogenation is a process where hydrogen directly combines with unsaturated organic compounds in the presence of catalyst like nickel.
- Oxidation reaction involves : the addition of oxygen, removal of hydrogen, addition of electronegative ion, removal of electropositive ionfrom a substance.
- Reduction reaction involves- the removal of oxygen/addition of hydrogen/addition of electropositive ion/removal of electronegative ion from a substance.
- In some chemical reactions, oxidation and reduction occur together, where one substance gets oxidised while other gets reduced. Such reactions are called redox reactions.

EXERCISE

A. VERY SHORT ANSWER TYPE QUESTIONS

1. Give one point of similarity between Hydrogen and Alkali metals.
2. Give one point of similarity between Hydrogen and Halogens.
3. Give the following preparation of Hydrogen with :
   (a) Sodium (b) Potassium
4. What is Activity series ?
5. Give two reasons, why zinc is the most feasible metal to prepare hydrogen.
6. Give only the reactions involved in the preparation of hydrogen in the lab :
   (a) By Bosch Process
   (b) By Electrolysis of Water.
7. Give one chemical property of hydrogen.
8. Give two uses of Hydrogen.
9. What is oxidation, Reduction and Redox reaction ?

B. SHORT ANSWER TYPE QUESTIONS

1. Explain each of the following statements :
   (a) Reaction of aluminium with water is comparatively slow and gradually reaction stops.
   (b) Amongst all the metals, zinc is preferred to prepare hydrogen gas.
   (c) A burning splinter brought near the mouth of test tube containing hydrogen gas extinguished with a popping sound.
   (d) Hydrogen is collected by downward displacement of water.
   (e) Sodium and potassium are not preferred as the starting metals in the preparation of hydrogen gas from cold water.
   (f) Sodium and potassium metals are stored in kerosene.
2. Write any two examples of reverisible reaction.

C. LONG ANSWER TYPE QUESTIONS

(i) Explain each of the following statements :

1. Hydrogen gas is used for producing artificial petrol from coal.
2. Hydrogen gas is used in the extraction of metals.
3. Yellow solid deposit is found when hydrogen sulphide reacts with an oxidising agent.
4. Distilled water does not support the process of electrolysis, thus acidified water is used.

(ii) Explain the role of each of the following chemical reagents

1. Caustic potash solution in Bosch process
2. Ammoniacal cuprous chloride solution in Bosch process
3. Silver nitrate in the purification of hydrogen gas
4. Lead nitrate in the purification of hydrogen gas
5. Caustic potash solution in the purification of hydrogen gas

(iii) Explain each of these terms by giving an example.

1. Hydrogenation of oil
2. Equilibrium
3. Redox reaction

(iv) Explain the process of electrolysis of water with a neat labelled diagram.

D. FILL IN THE BLANKS :

(i)

1. Relative molecular mass of hydrogen is ....
2. When CuO reacts with magnesium, ... is oxidised and .... is reduced.
3. Colour of acidified potassium dichromate changes from .... to .... when reacted with a reducing reagent.

4. During electrolysis, oxygen is collected at ....and hydrogen is collected at ....
5. Calcium hydride reacts with water to form .... and hydrogen.
6. .... is the most readily oxidised metal in the activity series.
7. Hydrogen in the form of .... flame is used for welding and cutting of metals.
8. Conversion of vegetable oil into vanaspati ghee is carried out by the process of ....
9. Combination of carbon monoxide and Hydrogen is called as ....
10. When Potassium metal catches fire, it burns with a .... colour flame.

(ii) Complete and balance following equations :

Metal	Acid/alkali	Salt	Hydrogen
Mg^+	$\rightarrow$		H_2
K^+	$\rightarrow$		H_2
Na^+	$\rightarrow$		H_2
Fe^+	$\rightarrow$		H_2
Zn^+	$\rightarrow$		H_2
Pb^+	$\rightarrow$		H_2

E. TRUE OR FALSE

1. Hydrogen is more reactive than copper.
2. Water is obtained when hydrogen burns in air.
3. There is no similarity between halogens and Hydrogen.
4. ($CO + H_2$) is known as Gobar or Marsh Gas.
5. Oxidation is loss of oxygen atom.
6. Reduction is gain of electron.

[Ans : (1) True, (2) True, (3) False, (4) False, (5) False, (6) True.]

F. MATCH THE COLUMN

Column A	Column B
1. Cu	(a) ($CO + H_2$)
2. Al	(b) 1
3. Valency of Hydrogen	(c) Less reactive than hydrogen
3. Water gas	(d) Gain of electron
4. Oxidation	(e) More Reactive than hydrogen
5. Reduction	(f) Loss of electron

[Ans : 1. (c), 2. (e), 3. (b), 4. (a), 5. (f), 6. (d)]

G. MULTIPLE CHOICE QUESTIONS

1. Hydrogen is a .... gas.
(a) Acidic (b) Basic
(c) Neutral (d) Amphoteric

2. Valency of Hydrogen is :
(a) 0 (b) 3
(c) 2 (d) 1

3. Hydrogen burns in air to form :
(a) O_2 (b) H_2O_3
(c) H_2O (d) H_2O_2

4. Which metal among the following is least reactive than hydrogen ?
(a) K (b) Ca
(c) Na (d) Au

5. Method used for the laboratory preparation of Hydrogen is :
(a) Bosch process
(b) Ammonification process
(c) Claisen method
(d) Haber's process

[Ans : 1. (c), 2. (d), 3. (c), 4. (d), 5. (a)]

H. IDENTIFY THE TYPE OF REACTION (OXIDATION/REDUCTION/REDOX) AND JUSTIFY YOUR ANSWER

1. $ZnO + C \longrightarrow Zn + CO$

2. $MnO_2 + HCl \longrightarrow MnCl_2 + H_2O + Cl_2$

3. $Cl_2 + 2NaBr \longrightarrow 2NaCl + Br_2$

4. $3MnO_2 + 4Al \longrightarrow 3Mn + Al_2O_3$

5. $3C + 2Fe_2O_3 \longrightarrow 4Fe + 3CO_2$

6. $Fe + O_2 + H_2O \longrightarrow Fe_2O_3.\,xH_2O$

7. $HCl + KOH \longrightarrow KCl + H_2O$

✦ ✦ ✦

CHAPTER 7

STUDY OF GAS LAWS

LEARNING OUTCOMES

7.1. Introduction
7.2. Physical Properties of Gas
7.3. Effect of pressure, Temperature and Volume
7.4. The Gas Laws

7.1. INTRODUCTION

Matter exists in three states *viz.* Solid, Liquid and Gas. Gaseous state is characterised by weak intermolecular forces of attraction which gives them freedom to spread throughout the available space. Gas has definite mass but not definite volume.

7.2. PHYSICAL PROPERTIES OF GAS

Kinetic molecular theory explains characteristics of gas as follows :

1. Gases are made up of tiny particles which are in constant motion at high velocity.
2. Gases do not have fixed volume or shape because the space between the particles is more and intermolecular forces of attraction are weak which gives them freedom to spread throughout the available space.
3. Constantly moving gas particles exert pressure in all directions, when they collide with each other and hit the sides of the container. At a given time, temperature and volume, same number of particles hit the walls of container. Therefore, gases exert same pressure in all directions.
4. Number of gas particles in a given volume is much less compared to that of solid and liquid particles. As there is more space between gas particles, less number of particles can accommodate in the volume compared to solid and liquid. Therefore, gases have low density.
5. As there is more space between gas particles and particles move at a higher velocity, rate of diffusion is also higher.
6. As there is more space between gas particles, gases can be easily compressed. When pressure is applied, particles come closer and the volume is reduced. Gases can be liquified by applying pressure and reducing temperature.
7. When the temperature of container containing gas is increased, the particles gain kinetic energy from heat and start moving at higher speed. This leads to expansion and increase in the volume. Similar change is observed when pressure is reduced.

7.3. EFFECT OF PRESSURE, TEMPERATURE AND VOLUME ON GASES

According to kinetic theory, particles of gas are in constant and random motion. Therefore, they possess kinetic energy. Change in temperature effects kinetic energy as temperature is directly proportional to the kinetic energy of particles. Theoretically, when temperature is zero on the Kelvin scale, the kinetic energy is zero, and the motion of molecule is stopped. The temperature at

which the molecular motion completely ceases on the Kelvin scale is called **absolute zero.**

Gas particles which are in constant motion, collide with each other and hit the walls of the container, thus exerting equal pressure in all directions.

With the increase in temperature or decrease in pressure, volume of gas increases. When the temperature is increased, the particles gain energy and start moving apart at a higher speed. Thus, increasing the volume. If this process happens in a closed non-expandable container, the pressure inside will increase. This is the basic principle of working of a pressure cooker. Similarly, when pressure is reduced, particles gain freedom and move apart, thus increasing the volume.

7.4. THE GAS LAWS

The gas laws describe the behaviour of gas under known conditions of pressure, volume and temperature.

Boyle's Law

It is the mathematical relationship between pressure and volume at constant temperature.

Consider the following experiment :

Procedure : Take a syringe fitted with a piston. Raise the piston and wrap the nozzle with an adhesive tape. Fit the nozzle in a rubber cork. Now, lower the piston by placing some weight on the piston. Record your observations. Now gradually remove the weights. Record your observations again.

Observation : As the piston is lowered, the volume of air decreases. Further when the weights are removed, volume of air increases.

Conclusion : At constant temperature, volume of the gas is inversely proportional to the pressure applied on the system. Increase in pressure decreases the volume of gas while, decrease in pressure increases the volume of gas.

Boyle's law states that volume of a given mass of dry gas is inversely proportional to the pressure exerted at constant temperature.

Mathematical Expression :

Suppose a gas occupies a volume V_1 when it is at pressure P_1, then

$$V_1 \propto \frac{1}{P_1}$$

or $$V_1 = \frac{k}{P_1}$$

or $$P_1V_1 = k = \text{Constant}$$

If V_2 is the volume occupies when the pressure is P_2 at the same temperature, then

$$V_2 \propto \frac{1}{P_2}$$

or $$V_2 = \frac{k}{P_2}$$

or $$P_2V_2 = k = \text{Constant}$$

$$P_1V_1 = P_2V_2$$

$$= k\text{; at constant temperature}$$

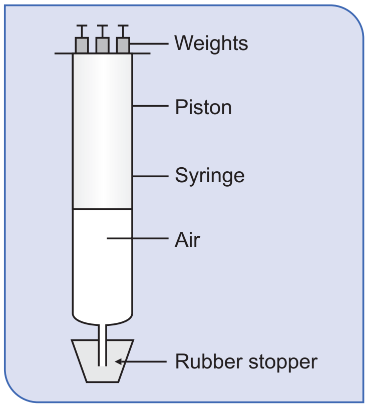


Fig. 7.1 : Pressure-Volume relationship

Graphical verification of Boyle's law

(a) V v/s $\frac{1}{P}$ gives a straight line passing through the origin.

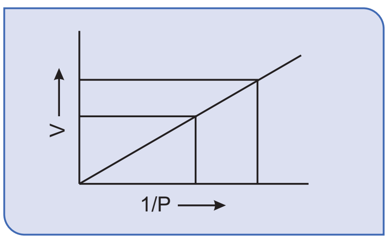


Fig. 7.2 : Straight line passing through origin

(b) V v/s P gives a hyperbolic curve

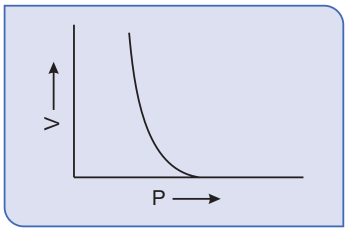
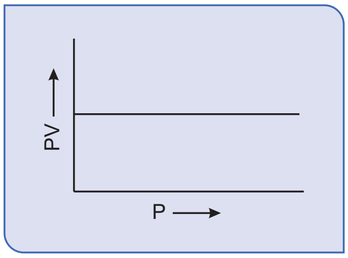


Fig. 7.3 : Hyperbolic Curve

(c) PV *v/s* P gives a straight parallel line to the X-axis



Significance of Boyle's law :

On increasing the pressure, the volume decreases and density increases. Air is less dense in high altitudes because atmospheric pressure is less. Therefore, it is difficult to breathe at higher altitudes as lesser amount of oxygen is available. Hence, mountaineers carry oxygen cylinders with them.

SOLVED NUMERICALS

1. A 3.20 L sample of gas has a pressure of 102 kPa. If the volume is reduced to 0.650 L, what pressure will the gas exert ?

Solution :

$$V_1 = 3.20 \text{ L}$$
$$P_1 = 102 \text{ kPa}$$
$$V_2 = 0.650 \text{ L}$$
$$P_2 = ?$$
$$P_2 = \frac{P_1 \times V_1}{V_2}$$
$$\frac{102 \times 3.2}{0.650} = 502 \text{ kPa}$$

2. A balloon is filled with hydrogen at room temperature. It will burst if pressure exceeds 0.2 bar. If at 1 bar the gas occupies 22.7 Volume, upto what volume, can the balloon be expanded ?

$$V_1 = 2.27 \text{ L}$$
$$P_1 = 1 \text{ bar}$$
$$V_2 = ?$$
$$P_2 = 0.2 \text{ bar}$$
$$V_2 = \frac{P_1 \times V_1}{P_2}$$
$$\frac{2 \times 2.27}{0.2} = 11.35 \text{ L}$$

Charles's Law

It is the mathematical relationship between temperature and volume at constant temperature.

Charles law states that volume of given mass of dry gas is directly proportional to its absolute Kelvin temperature, if pressure is kept constant.

As pressure remains constant, the volume of a given mass of dry gas increases or decreases to 1/273 of its volume at 0°C for every 1°C increase or decrease in the temperature.

According to this law, temperature is directly proportional to the volume. Increase in temperature leads to increase in volume and vice-versa

Mathematical expression of Charles law :

Suppose a gas occupied a volume V_1 when its temperature is T_1, then

$$V_1 \propto T_1$$

or $$V_1 = kT_1$$

or $$\frac{V_1}{T_1} = k = \text{Constant}$$

If V_2 is the volume occupied when the pressure is T_2 at the same pressure, then

$$V_2 \propto T_2$$

or $$V_2 = kT_2$$

or $$\frac{V_2}{T_2} = k = \text{constant}$$

$$V_1T_1 = \frac{V_2}{T_2}$$

$= k$ (at constant temperature)

Graphical representation of Charles law :

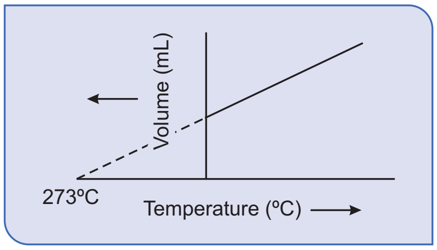


Fig. 7.4 : Graphical representation of Charles law

A Fig 7.4 shows the relationship between volume and temperature, where a straight line is observed.

Siginificance of Charles law : Since the volume of a given mass of a gas is directly proportional to its temperature, hence the density decreases with increase in temperature. This is the reason why hot air is filled in balloons used for meteorological purposes.

Absolute zero :

According to Charles law,

V_0 = the volume of gas of fixed mass at 0°C and

V = the volume the same gas at temperature *t* can be expressed as follows :

$$V = V_0\left(\frac{273 + t}{273}\right)$$

Now, volume at – 273°C will be

$$V = V_0\left(\frac{273 - 273}{273}\right) = 0$$

Above equation shows that the volume of gas reduces to zero at temperature – 273°C. The temperature – 273°C is called absolute zero. All gases will condense and liquify before reaching – 273°C. Theoretically this is the lowest temperature, that any system can attain. However it is not possible to attain this temperature practically.

SOLVED EXAMPLES

1. A 600 mL sample of nitrogen is heated from 27°C to 77°C at constant pressure. What is the final volume ?

Solution :

V_1 = 600 mL

T_1 = 27 °C

V_2 = ?

T_2 = 77 °C

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

$$V_2 = \frac{27 \times 600}{77} = 210.39 \text{ ml}$$

2. Carbon dioxide is usually formed when gasoline is burned. If 30.0 L of CO_2 is produced at a temperature of 1.00×10^3 °C and allowed to reach room temperature (25.0 °C) without any pressure changes, what is the new volume of the carbon dioxide ?

Solution :

V_1 = 30 L

T_1 = 1000 °C = 1273 K

V_2 = ?

T_2 = 25 °C = 298K

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

$$V_2 = \frac{1273 \times 30}{298} = 128{\cdot}15 \text{ K}$$

The Gas Equation

The gas equation is the combination of Boyle's law and Charles law.

According to Boyle's law, $V \propto \frac{1}{P}$

According to Charles law, $V \propto T$

If we combine both the equations,

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

$$V = \frac{T}{P} \propto \text{constant or } \frac{PV}{T} = \text{Constant}$$

For given volume which changes from V_1 to V_2 at changing temperature T_1 to T_2

$$\boxed{\frac{P_1V_1}{T_2} = \frac{P_2V_2}{T_2}}$$

Standard Temperature and Pressure

STP is often used for measuring gas density and volume. Gases are prone to changing volumes with changes in pressure or temperature. Therefore,

there is need for a standard unit and value to measure temperature, volume and pressure of a given gas. Standard temperature is equal to 0°C, which is 273.15 K. Standard Pressure is 1 Atm, 101.3 kPa or 760 mm Hg or torr. At STP 1 mole of any gas occupies 22.4L.

SOLVED NUMERICAL

1. A child has a toy balloon with a volume of 1.80 litres. The temperature of the balloon when it was filled was 20°C and the pressure was 1.00 atm. If the child were to let go of the balloon and it rose 3 kilometres into the sky where the pressure is 0.667 atm and the temperature is – 10°C, what would the new volume of the balloon be?

$$V_1 = 1.8 \text{ L}$$
$$P_1 = 1 \text{ atm}$$
$$T_1 = 20°C = 293 \text{ K}$$
$$P_2 = 0.067 \text{ atm}$$
$$T_2 = -10°C = 263\text{K}$$
$$V_2 = ?$$
$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$
$$V_2 = \frac{P_1V_1T_2}{P_2T_1}$$
$$V_2 = \frac{1 \times 1.8 \times 293}{0.067 \times 263} = 29.93 \text{ L}$$

2. A small research submarine with a volume of 1.2×10^5 L has an internal pressure of 1.0 atm and an internal temperature of 15°C. If the submarine descends to a depth where the pressure is 150 atm and the temperature is 3°C, what will be the volume of the gas inside be if the hull of the submarine breaks?

$$V_1 = 120000 \text{ L}$$
$$P_1 = 1 \text{ atm}$$
$$T_1 = 15°C = 288\text{K}$$
$$P_2 = 150 \text{ atm}$$
$$T_2 = 3°C = 279 \text{ K}$$
$$V_2 = ?$$
$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$
$$V_2 = \frac{P_1V_1T_2}{P_2T_1}$$
$$V_2 = \frac{1 \times 120000 \times 288}{150 \times 279} = 825.8 \text{ L}$$

Scales of Measurement

Three different scales are commonly used to measure temperature : Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K). Thermometers measure temperature by using materials that expand or contract when heated or cooled. Mercury or alcohol thermometers, for example, have a reservoir of liquid that expands when heated and contracts when cooled, so the liquid column lengthens or shortens as the temperature of the liquid changes.

(i) Celsius scale of temperature : This scale is based on the melting and boiling points of water under normal atmospheric conditions. The current scale is an inverted form of the original scale, which was divided into 100 increments. Because of these 100 divisions, the Celsius scale is also called the centigrade scale. In Celsius scale the boiling point of water is designated as 100°C and melting point of ice as 0°C.

(ii) Kelvin scale of temperature : The Kelvin is the primary unit of temperature measurement. It is defined as the temperature scale with its zero of 273·15 °C, whose each degree is equal to one degree on the cersius scale. The temperature 'T' in degree celsius is equal to T – 273·16.

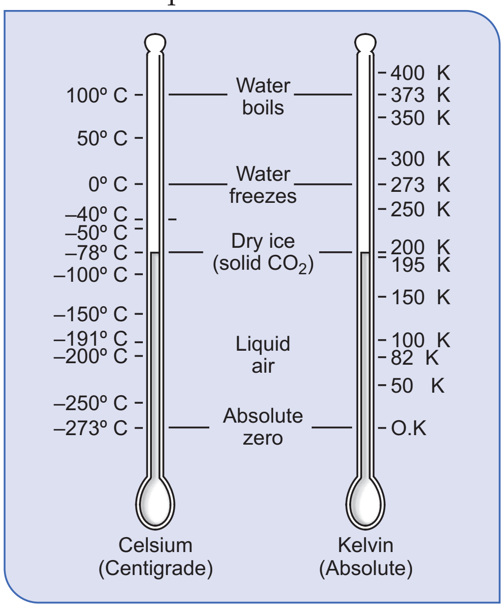


Fig. 7.5 : The two scales of Temperature.

Thus the two scales of temperature measurement are mathematically related as :

$$\boxed{\begin{array}{l} \textbf{T (in °C) + 273.15 = T (in K)} \\ \textbf{T (in K) – 273.15 = T (in °C)} \end{array}}$$

Also there exist another scale of temperature measurement known as **Fahrenheit Scale.** It is mathematically related to above two scales of temperature measurement as:

$$°C = \left(\frac{5}{9}\right) \times (°F) - 32$$

$$°F = \left(\frac{9}{5}\right) \times (°C) + 32$$

SOLVED EXAMPLES

1. Convert the temperature of 45°C to the Kelvin Scale.

Solution : Temperature on Celsius scale = 45°C

From the formula :

$$T(\text{in}°C) + 273.15 = T(\text{in K})$$
$$45 + 273.15 = T$$
$$\therefore \quad T = 318.15 \text{ K}$$

2. Convert the temperature 340 K to the celsius scale.

Solution : Temperature on Kelvin scale = 340 K

From the formula :

$$T(\text{in K}) - 273.15 = T(\text{in }°C)$$
$$340 - 273.15 = T$$
$$\therefore \quad T = 66.85°C$$

3. Convert the temperature of 40°C to the Kelvin and fahrenheit scale.

Solution : Temperature on Celsius scale = 40°C

From the formula :

$$T(\text{in }°C) + 273.15 = T(\text{in K})$$
$$40 + 273.15 = T(\text{in K})$$
$$T = 313.15 \text{ K}$$

Again, we have,

$$T(\text{in }°F) = (9/5) \times [T°C + 32]$$
$$= \frac{9}{5} \times [40 + 32]$$
$$= \frac{9}{5} \times 72 = 129{\cdot}6\ °F$$

SUMMARY

- Matter exists in three states *viz.* Solid, Liquid and Gas. Gaseous state is characterised by weak intermolecular forces of attraction which gives them freedom to spread throughout the available space.
- According to kinetic theory, particles of gas are in constant and random motion. Therefore, they possess kinetic energy. Change in temperature effects kinetic energy as temperature is directly proportional to the kinetic energy of particles.
- The temperature at which the molecular motion completely ceases on the Kelvin scale is called absolute zero.
- Gas particles which are in constant motion, collide with each other and hit the walls of the container, thus exerting equal pressure in all directions.
- Boyle's law states that volume of a given mass of dry gas is inversely proportional to the pressure exerted at constant temperature.
- Charles law states that volume of given mass of dry gas is directly proportional to its absolute Kelvin temperature, if pressure is kept constant.
- Standard temperature is equal to 0 °C, which is 273.15 K. Standard pressure is 1 Atm, 101.3 kPa or 760 mmHg or torr. At STP 1 mole of any gas occupies 22.4L.
- The gas equation is the combination of Boyle's law and Charles law. For given volume which changes from V_1 to V_2 at changing temperature T_1 to T_2

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

EXERCISE

A· VERY SHORT ANSWER TYPE QUESTIONS

1. Define the following term :
   (a) Matter (b) Diffusion
   (c) Absolute zero
2. Give reason for following observations :
   (a) Gases do not have fixed volume or shape
   (b) Constantly moving gas particles exert pressure in all directions.
   (c) Gases have lower density as compared to solids.
   (d) Rate of diffusion is higher in gases.
   (e) Gases are easily compressible.
3. Show the graphical representation of Boyle's law and Charles law.
4. The gas law is a combination of Boyle's law and Charles law. Justify the statement.
5. Give two characteristics of gases.
6. Convert :
   (a) 213 K to Celsius
   (b) 56°C to K
7. Give the mathematical form of the following :
   (a) Gas equation
   (b) Charles law
   (c) Boyle's law
8. Name the three scales of temperature measurement.

B· SHORT ANSWER TYPE QUESTIONS

1. What are the main postulates of the kinetic theory of matter ?
2. State how the kinetic theory of gases justifies the properties of an ideal gas.
3. State the properties of gases.
4. State Boyle's and Charles laws of gases.
5. What do you mean by N.T.P and S.T.P ?

C· LONG ANSWER TYPE QUESTIONS

1. Derive gas equation with the help of Boyle's and Charles' law.
2. Explain Boyle's and Charles' law with respect to kinetic theory of matter.

D· FILL IN THE BLANKS

1. As the volume of confined gas decreases at constant temperature, the pressure exerted by the gas ....
2. Kinetic theory is the concept that all matter is composed of particles and that they are in constant ....
3. Gas pressure is measurement of the number of ....between particles.
4. .... is a measure of the kinetic energy or speed of the particles in a substance.
5. When volume is doubled, pressure is .... by $\frac{1}{2}$.
6. When temperature of a gas filled balloon is doubled, the volume of the balloon is
7. When temperature of a canister is doubled, pressure is ....
8. The escape of gas through a small hole in a container is called ....
9. The pressure exerted by gas in a mixture is called........................ pressure.
10. When the gas is compressed, the temperature of the gas ....

[**Ans. :** 1. increases, 2. motion, 3. collisions, 4. temperature, 5. reduces, 6. doubled, 7. doubled, 8. effusion, 9. partial, 10. increases.]

E· TRUE OR FALSE

1. According to Charles law, if you have a balloon inside a car at noon during a hot summer day the balloon molecules inside will increase in pressure.
2. A good example of Charles law is when a piece of metal expands in the heat.
3. You drove continuously from Laguna to Manila and you observed that the pressure in your tires increased. This is because of the increased temperature outside the tire caused by friction.
4. The number of particles in a mole of a pure substance is 6.02 × 1023.
5. Pressure is a direct result of collisions between gas particles and the walls of their container.

6. According to Boyle's law, the volume of a gas is inversely proportional to its pressure at constant temperature.
7. The combined gas law states the relationship among pressure, temperature, and volume of a fixed amount of gas.
8. According to Gay-Lussac's law, pressure is directly proportional to temperature at constant volume.
9. According to the combined gas law, pressure is inversely proportional to volume and directly proportional to temperature, and volume is directly proportional to temperature.
10. Zero on the Celsius scale is also known as absolute zero.

[**Ans.** : 1. False, 2. True, 3. False, 4. True, 5. True, 6. False, 7. True, 8. True, 9. True, 10. False.]

F· MATCH THE COLUMN

Column A	Column B
1. Vacuum	(a) Result in pressure exerted by a gas
2. Kinetic energy	(b) A space where no particles of matter exist
3. Gas pressure	(c) The energy an object has because of its motion
4. Atmospheric pressure	(d) Results from the force exerted by a gas per unit surface area of an object
5. Collisions of particles	(e) Results from the collision of atoms and molecules in air with objects

[**Ans. :** 1. (b), 2. (c), 3. (d), 4. (e), 5. (a).]

G· MULTIPLE CHOICE QUESTIONS

1. General gas constant is represented by :
   (a) B (b) S
   (c) K (d) R
2. R is known as :
   (a) molar gas constant
   (b) volume constant
   (c) pressure constant
   (d) temperature constant
3. In Avogadro's law volume is directly propotional to :
   (a) Pressure (b) (no of moles)
   (c) Volume (d) Temperature
4. Laws of gases are obtained by :
   (a) pressure
   (b) volume
   (c) ideal gas equation
   (d) Joule Thomson effect
5. Uniform behaviour is shown by gases towards :
   (a) internal conditions
   (b) atmosphere
   (c) external conditions
   (d) sea level
6. If a nitrogen gas occupies a volume of 500 ml at a pressure of 0.971 atm. What volume will the gas occupy at a pressure of 1.50 atm, assuming the temperature remains constant ?
   (a) 342 mL (b) 424 mL
   (c) 324 mL (d) 442 mL
7. Not all Gas law problems have Kelvin (K) as the unit of temperature. They can be expressed in Celsius (°C) and Fahrenheit (°F). So convert 123°C to K.
   (a) 396 K (b) 486 K
   (c) 369 K (d) 458 K
8. Gas pressure is caused by :
   (a) Gas molecules heating up
   (b) Gas molecules reacting with other gas molecules
   (c) Gas molecules hitting the walls of a container
   (d) Gas molecules hitting other gas molecules
9. At a pressure of 5.0 atmospheres, a sample of gas occupies 40. liters. What volume will the same sample occupy at 1.0 atmosphere?
   (a) 0.0050 L (b) 0.13 L
   (c) 200 L (d) 8.0 L

10. Assuming that the temperature remains constant. How can you increase the pressure of a gas ?
    (a) Increase the container volume
    (b) Add more molecules of the gas
    (c) Decreases the container volume
    (d) None of the above

[**Ans. :** 1. (d) 2. (a) 3. (b) 4. (d) 5. (c) 6. (c) 7. (a) 8. (c) 9. (c) 10. (c).]

H· NUMERICAL PROBLEMS

1. A gas occupies 3 litres at 0°C. What volume will it occupy at –20°C, the pressure remaining constant ?
2. Determine the volume occupied by 2.34 grams of carbon dioxide gas at STP.
3. A sample of argon gas at STP occupies 56.2 litres. Determine the number of moles of argon and the mass in the sample.
4. A gas is allowed to expand at constant temperature from its initial volume of 400 mL to a final volume of 2000 mL. The final pressure is calculated to be 5 atm. Calculate the initial pressure.
5. At what temperature will 0.654 moles of neon gas occupy 12.30 litres at 1.95 atmospheres ?
6. A given mass of sulphur dioxide occupies 500 mL at 47°C. To what temperature must the gas be heated to change the volume to 650 mL? Assume that the pressure of the gas remains constant.
7. A 12.0 g sample of gas occupies 19.2 L at STP. What is the molecular weight of this gas ?
8. A gas occupies 500 mL at S.T.P. Find the volume occupied by the gas if the pressure is changed to 300 mm of Hg and the temperature to 17°C.
9. 892 c.c. of a gas is stored at -50°C and 540 mm of Hg pressure. Calculate the volume if the temperature increases by 57°C and the pressure increases by 20%.
10. A 600.0 mL sample of nitrogen is warmed from 77.0 °C to 86.0 °C. Find its new volume if the pressure remains constant.
11. A gas syringe contains 42.3 milliliters of a gas at 98.15 °C. Determine the volume that the gas will occupy if the temperature is decreased to – 18.50 °C.
12. If 15.0 liters of neon at 25.0 °C is allowed to expand to 45.0 liters, what must the new temperature be to maintain constant pressure ?
13. A sample of helium has a volume of 521 dm^3 at a pressure of 75 cm Hg and a temperature of 18° C. When the temperature is increased to 23°C, what is the volume of the helium ?
14. If a calm person with a lung capacity of 3.5 liters and a body temperature of 36°C gets angry, what will the volume of the person's lungs be if their temperature rises to 39°C ?
15. Convert the temperature of 50 °C to the Kelvin scale of temperature.
16. Convert the temperature of 310 K to the Celsius scale of temperature.
17. Convert the temperature of 48 °C to the :
    (a) Kelvin scale
    (b) Fahrenheit scale

✦ ✦ ✦

CHAPTER 8

ATMOSPHERIC POLLUTION

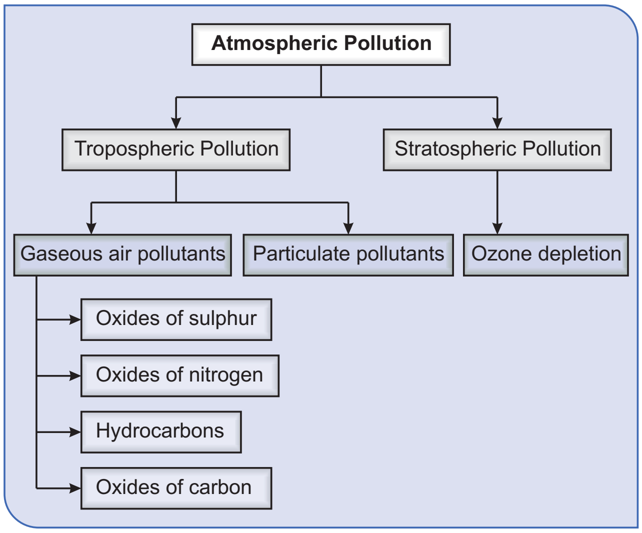
LEARNING OUTCOMES

- 8.1. Introduction
- 8.2. Air Pollution and Air Pollutants
- 8.3. Pollutant
- 8.4. Gaseous Air Pollutants
- 8.5. Acid Rain
- 8.6. Green House Effect
- 8.7. Mechanism of Global Warming
- 8.8. Ozone Depletion

8.1. INTRODUCTION

The word environment is descended from the Middle French preposition ***environ*** "around", ***environment,*** in its most basic meaning, is "that which surrounds." Environment means sum of all that conditions and elements which affects, directly or indirectly the life and progress of living organisms. In our surroundings, environment is made up of physical and biological elements. Environmental pollution is defined as the undesirable change in physical, chemical and biological characteristics of our air, land and water. Any undesirable introduction of any substance to air, water and soil by any natural source or due to human activity to a level of concentration that causes instability, disorder, harm or discomfort to the ecosystem *i.e.* physical systems or living organisms. The main causes of pollution are Fast population growth, Rapid urbanization, Excessive industrialisation and use of pesticides in agriculture.

Any substance, as certain chemicals or waste products that have undesirable effects, or adversely affect the usefulness of a resource is called as pollutant. It may have long or short term effects either by changing the rate of growth of living organisms or by interfering with human amenities, comfort and health or property values. Atmospheric pollution can be classified as :



8.2. AIR POLLUTION AND AIR POLLUTANTS

Air pollution worldwide is a growing threat to human health and the natural environment. Air pollution can be defined as, "the introduction of pollutants, organic molecules, or other unsafe materials into Earth's atmosphere, which can cause harm to human beings, other living entities such as animals and crops, or the natural or built environment". Air pollution may come from man-made or natural causes.

Air pollution results from a variety of causes, not all of which are within human control. Some of them are manmade whereas some are natural sources :

(a) Man-made sources : Man-made sources include combustion of fuels, use of auto mobiles, aquatic vessels, and planes. Deforestation, Fumes from hair spray, paint, aerosol sprays, varnish and other solvents, Military resources, such as nuclear weapons, toxic gasses, germ warfare and rocketry etc.

(b) Natural sources : It includes :

(i) Dust storms in desert areas and smoke from forest fires and grass fires contribute to chemical and particulate pollution of the air.

(ii) The most important natural source of air pollution is volcanic activity, which at times pours great amounts of ash and toxic fumes into the atmosphere.

8.3. POLLUTANT

Any substance, as certain chemicals or waste products that have undesirable effects, or adversely affect the usefulness of a resource is called as pollutant. It may have long or short term effects either by changing the rate of growth of living organisms or by interfering with human amenities, comfort and health or property values.

Types of Pollutants

(a) Primary pollutants : These are the pollutants which enter the atmosphere after their formation and do not undergo any chemical change. **Example:** SO_2, NO_2 etc.

(b) Secondary pollutants : These are the pollutants which are produced as a result of reaction between primary pollutants present or hydrogen present in the atmosphere. **Example :** H_2SO_4, SO_3 etc.

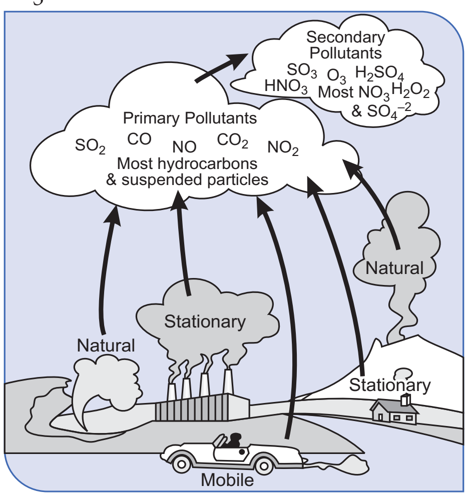


Fig. 8.1 : Types of pollutants

(i) Biodegradable pollutants : These are the pollutants which can be easily degraded and can be easily decomposed by microorganism in due period are called biodegradable pollutant. **Example :** Clothes, dead bodies, domestic sewage etc.

Fig. 8.2 : Biodegradable waste

(ii) Non-biodegradables pollutants : These are the pollutants which cannot be easily decomposed or degraded and their presence even in small quantity in the atmosphere is harmful to living beings. **Example :** mercury, aluminium, DDT etc.

Fig. 8.3 : Non- Biodegradable waste

Troposphere extends from 8 to 12 km above the earth's surface. About 80% of the total mass of the air and almost all of the water vapours of the atmosphere is found in troposphere. Troposphere pollution can be studied as gaseous air pollutants and particulate pollutants.

8.4. GASEOUS AIR POLLUTANTS

1. Oxides of sulphur : Sulphur dioxide is the main pollutant among sulphur oxides. The main sources are Volcanic eruption, Thermal power plants, Metallurgical processes, Combustion of

coal and fossil fuels and Oxidation of hydrogen sulphide. It can cause breathlessness, asthma, bronchitis and emphysema in human beings.

2. Oxides of nitrogen : NO and NO_2 are main pollutants among nitrogen oxide. The main sources are natural bacteria in soil, Combustion of automobile engines and thermal generating stations, and industrial processes. It can cause acid rain, contributes to global warming, hampers the growth of plants, NOx can form with other pollutants to form toxic chemicals.

3. Hydrocarbons : The main sources are trees, Domestic animals, Burning of fuels like coal, wood, kerosene oil etc., automobiles and industrial operations. It has carcinogenic effects on lungs and causes cancer, smog formation and ageing in plants and shedding of leaves, flowers and twigs.

4. Oxides of carbon :

(a) Carbon monoxide (CO) : The main sources are incomplete combustion of carbon, fossil fuels or carbon containing compounds, Forest fire or agricultural waste burning, Automobile engines and Bacterial decay of living or organic matter. It is very toxic in nature due to its ability to combine with hemoglobin and obstructs the normal transport of oxygen in the blood stream.

(b) Carbon dioxide (CO_2) : It is another oxide of carbon which is very useful yet a major contributor to global increase in the temperature of earth known as the ***Global warming***. Carbon dioxide is also produced by biological decay of plants. Carbon dioxide gas is also emitted during volcanic eruptions, decomposition etc.

8.5. ACID RAIN

Acid rain as the name suggests is the precipitation of acid in the form of rain. Carbon dioxide which is released as a byproduct of various natural process, reacts with tiny droplets of water vapours in the atmosphere, forming carbonic acid. This carbonic acid dissolves with rain water and precipitates down to the earth. This is natural acid rain.

$$CO_2 + H_2O \longrightarrow H_2CO_3$$

However, as a result of air pollution sulphur dioxide and nitrogen dioxide is released. These gases react with water vapour to form sulphuric acid and nitric acid. These acid dissolve with rain water to form acid rain. Rainwater is usually slightly acidic, with pH level between 5 and 6 which is less than normal water.

Formation of sulphuric acid and nitric acid

The gases that cause acid rain are sulphur dioxide (SO_2), nitrogen dioxide (NO_2) and to a lesser extent, carbon dioxide (CO_2)

Formation of nitrogen Dioxide :

$$N_{2\,(g)} + 2O_{2\,(g)} \rightarrow 2NO_{2\,(g)}$$

Both sulphur dioxide and nitrogen dioxide are acidic oxides and react with water to form acids.

Sulphur dioxide reacts with water to form sulphurous acid.

$$SO_{2\,(g)} + H_2O_{(l)} \rightarrow H_2SO_{3\,(aq)}$$

Substances in the upper atmosphere then catalyze the reaction between sulphurous acid and oxygen to form sulphuric acid.

$$2H_2SO_{3\,(aq)} + O_{2\,(g)} \rightarrow 2H_2SO_{4\,(aq)}$$

Similarly, nitrogen dioxide reacts with water to form a mixture of nitric acid and nitrous acid.

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

Substances in the atmosphere then catalyze the reaction between nitrous acid and oxygen causing the formation of more nitric acid.

$$2HNO_{2\,(aq)} + O_{2\,(g)} \rightarrow 2HNO_{3\,(aq)}$$

Both sulphuric acid and nitric acid are soluble in water and are the major acids present in acid rain. As this forms and falls onto the Earth's surface, these strong acids are also brought to the surface causing harmful effects on the built and the natural environment.

Causes

The main cause of acid rain is the presence of air pollutants, like sulphur dioxide and nitrogen oxides. They produce acids when combined with water. Acid rain is considered as the wet deposits of air pollutants, where it is combined with moisture before falling into the ground while air pollutants that fall without combining with moisture is called dry deposits. Sulphur dioxide is the main pollutant among sulphur oxides. The main sources of sulphur are Volcanic eruptions, Thermal power plants, Metallurgical processes, Combustion of coal and fossil fuels and Oxidation of hydrogen sulphide.

NO and NO_2 are main pollutants among nitrogen oxides. The main sources are Natural bacteria in soil, Combustion of automobile engines, Thermal generating stations and Industrial processes.

Both sulphur dioxide and nitrogen dioxide undergo oxidation and then they react with water resulting in the formation of sulphuric acid and nitric acid respectively. Following reaction will clarify the acid formation reaction :

$$2SO_{2\,(g)} + O_{2\,(g)} + 2H_2O_{\,(l)} \rightarrow 2H_2SO_{4(aq)}$$
$$4NO_{2\,(g)} + O_{2\,(g)} + H_2O_{2\,(l)} \rightarrow 4HNO_{3\,(aq)}$$

Effects :

Acid rain have harmful effects on human and environment. Some of them are :

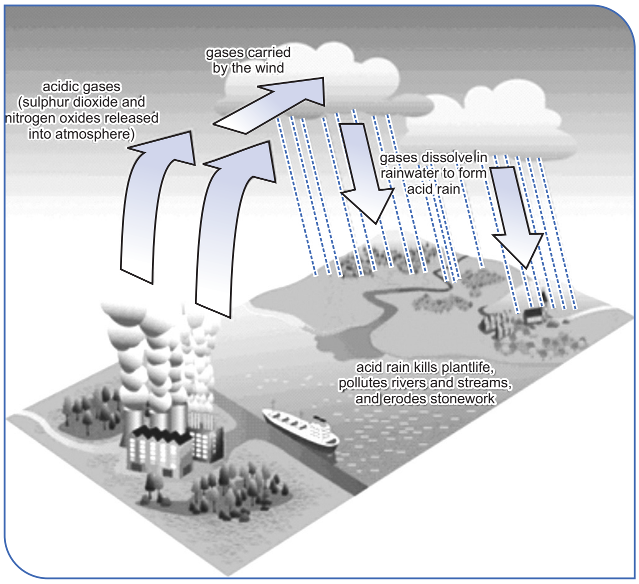


Fig. 8.4 : Effects of Acid rain

(i) It causes death of aquatic animals. Mixing of poisonous metals like copper, bronze in the water. When acid rain falls down and flows into the rivers and ponds it affects the aquatic ecosystem. As it alters the chemical composition of the water, to a form which is actually harmful for the aquatic ecosystem to survive and causes water pollution.

(ii) Also, acid rain wear away the waxy protective coating of leaves, damaging them and preventing them from being able to photosynthesise properly.

(iii) Acidification of soil and lakes which decreases the fertility of soil. Acid rain is very harmful for agriculture, plants and animals. It washes away all nutrients which are required for the growth and survival of plants. Acid rain affects agriculture by the way how it alters the composition of soil. It Deteriorates trees, forest and fields.

(iv) Lakes, rivers and marshes each have their own fragile ecosystem in which many different species of plants and animals are living which are dependent on one another for their survival. Suppose, if a species of fish disappears, the animals which feed on it will gradually disappear too and then the particular species of insect on which the extinct fish used to feed will start to grow, and their population increases. This in turn will affect the smaller insects or plankton on which the larger insect feeds. In this way, whole food chain is disturbed.

Remedial measures for stopping acid rain

Emission of oxides of sulphur and nitrogen should be reduced at all sources. Fuel containing low sulphur should be used. Scrubbers are installed at power stations to control sulphur emission to the atmosphere. In scrubber, water is sprayed to the sulphur dioxide. Thus, the gas rising gets absorbed and collected in the scrubber, which is later disposed. Solution of limewater can be also sprayed to sulphur dioxide. As a result of neutralization reaction between limewater and sulphur dioxide, a solid white coloured gypsum (calcium sulphate is formed).

INTEXT QUESTIONS

1. What are the sources of nitrogen oxides ?
2. Give two real life examples of acid rain.
3. What is the composition of acid rain ?
4. Give an example how acid rain affects aquatic food chain.

8.6. GREENHOUSE EFFECT

75% of the solar heat energy is absorbed by earth's surface. Remaining heat waves reflect back from the earth's surface. During this process some greenhouse gases like carbon dioxide, methane, CFCs etc. trap and absorb the heat waves. This process is known as global warming which increases the temperature of the earth's surface.

Greenhouse Gases

Carbon dioxide (CO_2)

Carbon dioxide (CO_2) is one of the greenhouse gases. It consists of one carbon atom with an oxygen atom bonded to each side. When its atoms are bonded tightly together, the carbon dioxide molecule can absorb infrared radiation and the molecule starts to vibrate. The main sources of CO_2 in atmosphere are decomposition, ocean release, natural gases, deforestation etc. Carbon dioxide is released into the air by respiration, Complete combustion of fossil fuels and carbon containing compounds, many industrial processes also produce a lot of carbon dioxide. Normally carbon dioxide is harmless and is not a pollutant. However, increasing concentration of CO_2 may affect the atmosphere causing undesirable change in climate. Eventually, the vibrating molecule will emit the radiation again, and it will likely be absorbed by yet another greenhouse gas molecule. This absorption-emission-absorption cycle serves to keep the heat near the surface, effectively insulating the surface from the cold of space.

Water vapour : Water Vapour is the most abundant greenhouse gas in the atmosphere, as the temperature of the atmosphere rises, more water is evaporated from ground storage (rivers, oceans, reservoirs, soil). Because the air is warmer, the absolute humidity can be higher (in essence, the air is able to 'hold' more water when it is warmer), leading to more water *vapour* in the atmosphere. As a greenhouse gas, the higher concentration of water vapour is then able to absorb more thermal IR energy radiated from the Earth, thus further warming the atmosphere. Water molecules capture the heat that earth radiates, and then re-radiates it in all directions, thus warming the earth's surface.

Methane (CH_4) : Methane is emitted during the production and transport of coal, natural gas and oil. Methane emissions also result from livestock and other agricultural practices and by the decay of organic waste in municipal solid waste landfills. It is a colourless, odourless, flammable gas, formed when plants decay in a region which is usually found near swamps and wetlands. Methane stays in the atmosphere for only 10 years, but traps 20 times more heat than carbon dioxide.

Nitrogen oxides :

Nitrogen monoxide is formed in high temperature combustion situations *e.g.* car engines, power station furnace burning coal, oil or natural gas.

$$N_{2\,(g)} + O_{2\,(g)} \longrightarrow 2NO_{(g)} \text{ (acidic gas)}$$

Nitrous oxide (N_2O)

In air the nitrogen monoxide rapidly combines with the oxygen in air

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

The nitrogen dioxide is oxidised to nitric acid by the reaction with oxygen from air when it dissolves in rainwater.

$$4NO_{2\,(g\text{-}air)} + O_{2\,(g\text{-}air)} + 2H_2O_{(l\text{-}rain)} \longrightarrow 4HNO_{3\,(aq\text{-}rain)}$$

Nitrous oxide is emitted during agricultural and industrial activities, as well as during combustion of fossil fuels and solid waste. It is the major contributor to greenhouse effect, it is a colourless gas with a sweet odour also called "laughing gas" and is primarily used as an anesthetic to reduce pain. It is important to reduce nitrous oxide emissions, because the nitrous oxide emissions of today will still be trapped in the atmosphere 100 years from now.

Chlorofluorocarbons : Chlorofluorocarbons, commonly known as CFCs, are a group of man-made compounds containing chlorine, fluorine and carbon. They are not found anywhere in nature. The production of CFCs began in the 1930s for the purpose of refrigeration. When CFCs are

broken down by ultraviolet (UV) radiation from the Sun, releasing free chlorine atoms which cause significant ***ozone depletion***. CFCs damage the protective ozone layer in the atmosphere, and as a result, more harmful ultraviolet rays from the sun are able to reach the earth's surface. Thus it is one of the important contributor to the increase in earth's atmosphere.

8.7. MECHANISM OF GLOBAL WARMING

Carbon dioxide has the capacity to absorb heat radiations emitted from earth surface in the form of infra red radiations (which have longer wave length and have a heating effect) and thus increases temperature. This increase in global temperature (global warming) is mainly due to carbon dioxide concentration and is called Greenhouse effect. The green house effect was first described by the French mathematician **J. Fourier** in 1827. Due to heavy industrialization and transportation, carbondioxide concentration is increasing day by day in the atmosphere. Besides carbon dioxide, some other gases associated with green house effect are methane, oxides of nitrogen and ozone. These are called greenhouse gases. Carbon dioxide is the most important green house gas.

These gases never let the radiations escape from the earth and hence the surface temperature of the earth increases and this is called **Global warming.** Hence, the greenhouse effect leads to global warming.

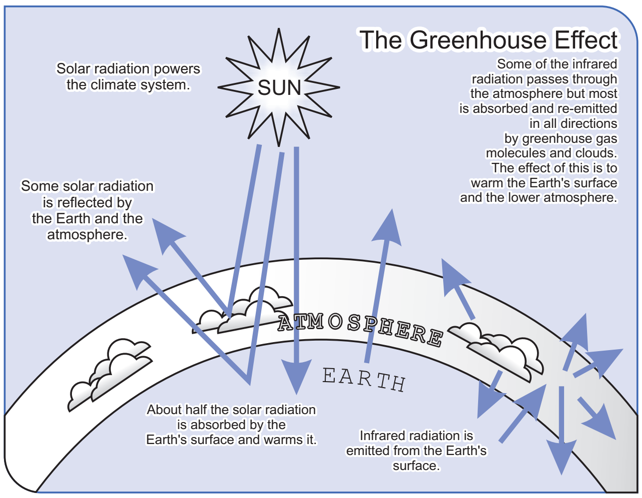


Fig. 8.5. Global Warming mechanism

Sources of Greenhouse gases : The main source of greenhouse gases is fossil fuel combustion. Some of the other sources are :

1. Electricity production : Electricity production generates the largest share of greenhouse gas emissions. Mostly, electricity is generated by burning of fossil fuels.

2. Transportation : This is also the cause of greenhouse gas emissions. It primarily comes from burning fossil fuel for our cars, trucks, ships, trains, and planes. Over 90 percent of the fuel used for transportation is petroleum based, which includes gasoline and diesel.

3. Industry : Greenhouse gas emissions from industry primarily come from burning fossil fuels for energy, as well as greenhouse gas emissions from certain chemical reactions necessary to produce goods from raw materials.

4. Agriculture : Excessive use of nitrogenous fertilizers in the agriculture also causes the concentration of these gases in the atmosphere to rise.

Ways of reducing greenhouse gases in the atmosphere :

- **(i) Reduction of use of fossil fuel :** By reducing the use of fossil fuels we can reduce the concentration of greenhouse gases in the atmosphere.
- **(ii) Improving energy efficiency :** If the energy efficiency of automobiles is reduced, it can reduce emission of greenhouse gases.
- **(iii) Reduction in deforestation :** Trees absorb carbondioxide from the atmosphere. If the amount of trees is decreased, concentration of carbon dioxide will increase in the atmosphere. Therefore, deforestation should be banned.
- **(iv) Afforestation :** More and more trees should be planted to reduce green house gases and thereby reducing the impact of global warming.
- **(v)** Reduction in the use of nitrogenous fertilizers.

INTEXT QUESTIONS

1. What are green house gases ?
2. How transportation increases the amount of green house gases in the atmosphere ?
3. What do you mean by global warming ?

8.8. OZONE DEPLETION

Formation of Ozone Layer

The region of atmosphere which is in contact of earth is troposphere. The layer next to troposphere is stratosphere. This area about 15-20 km above the sea is called Stratosphere. This layer contains nitrogen, oxygen, ozone and water vapour in small amount. The upper stratosphere consists of considerable amount of ozone (O_3). The ozone is a layer of oxygen (O_3) molecules in the Earth's atmosphere. The layer is present at an altitude of about 10-50 kilometre, with a maximum concentration in the stratosphere. The thickness of the ozone in a column of air from the ground to the top of the atmosphere is measured in terms of **Dobson Units** (DU). UV radiation from sun dissociates ozone into O_2 and O, but they quickly recombine to form O_3. In nature equilibrium is established between generation and destruction of O_3.

$$O_2 \longrightarrow O + O$$
$$O + O_2 \longrightarrow O_3$$
$$O_3 \longrightarrow O_2 + O$$

In the presence of ultraviolet radiations, Chloro-fluorocarbons are dissociated into nascent chlorine and ClO free radical. These intermediates are chemically active and convert ozone to oxygen. In this way, due to continuous formation of chlorine free radicals regularly dissociate ozone to oxygen.

$$O_{2\,(g)} \xrightarrow{uv} O_{(g)} + O_{(g)}$$
$$O_{(g)} + O_{2\,(g)} \xrightarrow{uv} O_3(g)$$
$$\dot{C}l_{(g)} + O_{3\,(g)} \longrightarrow Cl\dot{O}_{(g)} + O_{2\,(g)}$$
$$Cl\dot{O}_{(g)} + O_{(g)} \longrightarrow \dot{C}l_{(g)} + O_{2\,(g)}$$
$$CF_2Cl_{2\,(g)} \xrightarrow{uv} \dot{C}l_{(g)} + \dot{C}F_2Cl_{(g)}$$

Function of ozone in the atmosphere : The upper stratosphere consists of considerable amount of ozone (O_3), which protects us from the about 99% of harmful ultraviolet (UV) radiations (255 nm) coming from the Sun. The ozone layer naturally shields Earth's life from the harmful effects of the Sun's ultraviolet (UV) radiation. This protects the human beings and other organisms on the earth from the harmful effects of these radiations. Thus, it acts as a blanket in the atmosphere.

Destruction (Depletion) of ozone layer : In recent years, scientists have measured a seasonal thinning of the ozone layer primarily at the South Pole. This phenomenon is being called the ***ozone hole*** (ozone layer depletion).

Effects of ozone layer depletion :

(i) Damages DNA due to mutation

(ii) Ageing of skin

(iii) Damages skin and cause skin cancer

(iv) Inflammation of cornea **(Snow-blindness** cataract)

(v) Lethal to microbes

(vi) Immune system deficiency in humans.

(vii) UV-B radiation affects photosynthesis in plants.

(viii) It also damages nucleic acids in both aquatic and terrestrial vegetation.

The main reason of ozone layer depletion is believed to be the release of chlorofluorocarbon compounds (CFCs), also known as freons. The chlorofluorocarbon compounds are chemically inactive and do not dissolve in rain water. As a result, they reach stratosphere over a period of time. These chemicals react with ozone present in the Stratosphere and cause depletion of ozone. In the presence of ultraviolet radiations,

Chlorofluorocarbons are dissociated into nascent chlorine and ClO free radical. These intermediates are chemically active and convert ozone to oxygen. In this way, due to continuous formation of chlorine free radicals regularly dissociate ozone to oxygen.

INTEXT QUESTIONS

1. What is stratosphere ?
2. Which chemicals are responsible for ozone layer depletion ?
3. Write full form of CFC's.

Prevention Measures :

(a) **Reduction of use of fossil fuel :** We can use bicycles or walk to or from our schools or work places. Thus, Reducing the overall number of vehicles on the road, reducing the cumulative amount of cars resulting in pollution. We can also switch fuels from diesel or petrol to CNG or use unleaded petrol or diesel only

(b) **Improving energy efficiency :** We can use energy-efficient devices, such as CFL's and LED's rather than conventional light bulb's

(c) **Reduction in deforestation :** Ban on cutting of trees.

(d) **Afforestation :** More and more tress should be planted.

(e) Reduction in the use of nitrogenous fertilizers.

(f) We must switch off all lights, computers and all electrical appliances when not in use.

(g) Burning of fossil fuels is responsible for a major portion of air pollution. So need to develop and employ alternative non-combustive sources of energy. Some being nuclear power, geothermal power, solar power, tidal power, wind power, etc.

(h) Spreading awareness about the dangers of air pollution is something that all of us can definitely do.

SUMMARY

- Environmental pollution is defined as the undesirable change in physical, chemical and biological characteristics of our air, land and water.
- Acid rain, Global warming and ozone depletion are major hazards to the environment.
- Rainwater combines with oxides of nitrogen and sulphur present in the atmosphere and forms acid. When this falls as rain, it is known as acid rain.
- Acid rain have harmful effects on human beings, aquatic life and soil fertility.
- Carbon dioxide molecules trap a part of heat radiated by earth thereby maintaining the temperature of earth. It is called greenhouse effect.
- The increasing greenhouse effect due to increase in concentration of greenhouse gases is known as global warming.
- Carbon dioxide, methane, water vapour etc. are greenhouse gases.
- Reduction in use of fossil fuels, controlling deforestation, restricting the use of CFC's are some measures taken to reduce concentration of greenhouse gases in the atmosphere.
- Ozone protects us from harmful effects of ultraviolet rays.
- CFC's are responsible for depletion of ozone layer.

EXERCISE

A. VERY SHORT ANSWER TYPE QUESTIONS

1. Why greenhouse effect is important ?
2. Where have we discovered a hole in ozone layer ?
3. Which gas contributes largely in global warming ?
4. Name the gases which cause acid rain.
5. What is pollution ?
6. What is ozone ?
7. Which gas is present in large amount in upper part of stratosphere ?
8. Which chemicals are responsible for ozone depletion ?
9. Which fuel can be used in vehicles to reduce pollution ?
10. What chemicals are involved in global warming and ozone depletion ?
11. Why it is recommended to plant more trees ?

B. SHORT ANSWER TYPE QUESTIONS

1. What is an ozone layer ? Why is it useful for organisms on earth ?
2. When does acid rain occur ?
3. What are the effects of acid rain ?
4. What is the difference between stratospheric and ground layer ozone ?
5. Write the sources of greenhouse gases ?
6. Why CFC's are banned in some countries ?
7. Why it is mandatory to take a pollution certificate for vehicles ?
8. What is the pH of acid rain ?
9. What is the reason of corrosion of Taj Mahal ?

C. LONG ANSWER TYPE QUESTIONS

1. Is Ozone depletion and global warming inter-related? Explain.
2. What is acid rain? What are main sources of gases which are responsible for acid rain ?
3. What is the effect of ozone layer depletion ?

D. FILL IN THE BLANKS

1. The region of stratosphere is ....troposphere.
2. Acid rain ....the fertility of the soil.
3. Rain water combines nitrogen oxide in the atmosphere and form..........
4. Carbon dioxide gas can trap.........… radiations.
5. The gases which causes global warming are called ……….. gases.
6. ....gas is found in Stratosphere.
7. Rain water is more ....than normal.

[**Ans :** (1) above (2) reduces (3) nitric acid (4) infrared (5) greenhouse (6) Ozone (7) acidic]

E. TRUE OR FALSE

1. Acid rain is due to presence of oxides of only sulphur in air.
2. Mesosphere is higher than stratosphere.
3. Natural decay of animal and vegetable matter is the main source of NO.
4. The pH of rain water is above 7.
5. Carbon dioxide is the most abundant green-house gas.
6. Acid rain is the result of water pollution.

[**Ans :** (1) False, (2) True, (3) False, (4) False, (5) True, (6) False]

F. MATCH THE COLUMN

Column A	Column B
1. Oxides of sulphur	(a) Global warming
2. Carbon dioxide	(b) Carbon dioxide
3. Acid rain	(c) Chlorofluorocarbons
4. Depletion of ozone layer	(d) Respiratory disease

[**Ans :** 1.→(d), 2.→(a), 3.→(b), 4.→(c)]

G. MULTIPLE CHOICE QUESTIONS

1. Which of the following is **not** a greenhouse gas ?
   (a) CO_2 (b) CH_4
   (c) O_3 (d) N_2
2. What is the reason of increase in temperature ?
   (a) Acid rain
   (b) Nuclear accidents
   (c) Increase in heat at faster rate
   (d) Greenhouse effect

3. CFC's are used in :
(a) Refrigerants
(b) Plastic formation
(c) In electrical appliances
(d) All the above

4. In ozone, number of oxygen atoms are :
(a) 3 (b) 2
(c) 1 (d) 4

5. Acid rain is caused by increase in the atmospheric concentration of :
(a) Sulphur and nitrogen oxides
(b) Sulphur and carbon oxides
(c) Carbon and nitrogen oxides
(d) Water vapour

6. Greenhouse effect is due to the presence of :
(a) Ozone layer
(b) Moisture
(c) Carbon dioxide molecules
(d) Ultraviolet rays

7. The concentration of greenhouse gases is increasing because of :
(a) use of refrigerators
(b) increased combustion of fossil fuels
(c) deforestation
(d) all the above

8. Formation of ozone layer is maximum over :
(a) India (b) Antarctica
(c) Europe (d) Africa

9. The stratosphere ozone depletion leads to :
(a) Global warming
(b) Skin cancer
(c) Forest fires
(d) All the above

10. Taj Mahal is threatened due to the effect of :
(a) Chlorine (b) Sulphur dioxide
(c) Oxygen (d) Hydrogen

11. CFC's are not recommended to use in refrigerators because they :
(a) increase temperature
(b) depletes ozone
(c) affect environment
(d) affect human health

[**Ans :** 1. (d), 2. (b), 3. (a), 4. (a), 5. (a), 6. (c), 7. (d), 8. (b), 9. (d), 10. (b), 11. (b)]

✦ ✦ ✦

CHAPTER 9

PRACTICAL WORK

IDENTIFICATION OF GASES

S.No	Test	Observations	Colour and odour/ litmus test	Inference and Explanation
1	Apply a lit splint or ignite the sample.	A pop sound is heard.	Colourless and odourless, neutral to litmus	Hydrogen is present .
2	Take a glowing splint and place it in a sample of gas	It re-ignites the flame	Colourless and odourless, neutral to litmus	Oxygen is present
3	Bubble the gas into limewater (aqueous calcium hydroxide solution).	It turns cloudy–fine milky white precipitate of calcium carbonate	Colourless and odourless, changes moist blue litmus red	Carbon dioxide is present $Ca(OH)_{2\,(aq)} + CO_{2\,(g)} \rightarrow CaCO_{3\,(s)} + H_2O_{(l)}$
4	(i) Apply damp blue litmus paper. (Can use red litmus paper and see bleaching effect.) (ii) Dip a glass rod in silver nitrate solution and expose this glass rod to the gas.	(i) Blue litmus paper turns red and then is bleached white. (ii) White fumes are observed.	Greenish yellow colour, pungent odour. Turns blue litmus red which later turns colourless.	Chlorine is present Dense white fumes are observed due to the formation of Silver chloride.

S.No	Test	Observations	Colour and odour/ litmus test	Inference and Explanation
5	(i) Apply damp blue litmus paper. (ii) Dip a glass rod in silver nitrate solution and expose this glass rod to the gas.	(i) Litmuspaper turns red but does not bleach litmus paper. (ii) White fumes are observed.	Colourless and pungent smell, changes moist blue litmus paper red.	Hydrogen chloride is present (i) Strongly acidic gas. (ii) In water forms chloride ions hence precipitate with silver nitrate
6	Gas passed through a concentrated, acidified solution of potassium di-chromate ($K_2Cr_2O_7$).	Solution changes from orange to green.	Colourless, smells like burning Sulphur, changes moist blue litmus paper red and then bleaches it.	Sulphurdioxide is present The orange dichromate(VI) ion, $Cr_2O_7^{2-}{}_{(aq)}$ is reduced to the green $Cr^{3+}{}_{(aq)}$ ion.
7	Test gas with damp lead(II) ethanoate paper (lead acetate)	Rotten egg smell of hydrogen sulphide gas is obtained and the H_2S gas turns lead(II) ethanoate paper black.	Colourless, smells like rotten egg, changes moist blue litmus paper red	Hydrogen sulphide is present. Lead ethanoate reacts with hydrogen sulphide to give a black precipitate of Lead sulphide.
8	(i) Expose damp red litmus paper to the fumes. (ii) Dip a glass rod concentrated Hydrochloric acid solution and expose this glass rod to the gas.	(i) Red litmus paper turns blue. (ii) Gives white clouds with HCl fumes.	Colourless and pungent smell. Turns moist red litmus blue.	Ammonia (i) Ammonia is present. (ii) Ammonia reacts with hydrochloric acid to form ammonium chloride.
9	(i) Add a few drops to white anhydrous copper(II) sulphate. (ii) Dip in a piece of dry blue cobalt chloride paper.	(i) Turns from white to blue. (ii) Turns from blue to pink.	Colourless and odourless, cobalt chloride paper	Water is present

S.No	Test	Observations	Colour and odour/ litmus test	Inference and Explanation
10	Add a few drops of acidified ferrous sulphate solution.	Turns green solution to brown.	Reddish brown colour, irritating odour, turns blue litmus red.	Nitrogen dioxide gas is present.

Experiment No. 1

Action of heat on the following compounds :

(a) Copper carbonate, zinc carbonate

(b) Washing soda, copper sulphate crystals

(c) Zinc nitrate, copper nitrate and lead nitrate

(d) Ammonium chloride, iodine, ammonium dichromate.

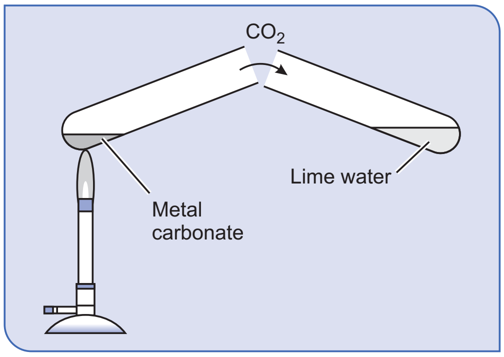
Answer : (a) Copper carbonate, zinc carbonate.

(i) Take two test tubes.

(ii) Place a small amount of copper carbonate in one test tube and zinc carbonate in the other.

(iii) Heat the solid gently and then strongly.

(iv) Test the gas with lime water.



Observation :

Carbonate tested	Colour of metal carbonate before heating	Compound formed and gas evolved if any	Decomposition easy or difficult	Colour of metal carbonate after heating
Copper carbonate	Bluish green	Copper oxide is formed and carbon dioxide is evolved.	Easy	White
Zinc carbonate	White	Zinc oxide is formed and carbon dioxide is evolved.	Easy	Yellow

(b) Washing soda and copper sulphate crystals

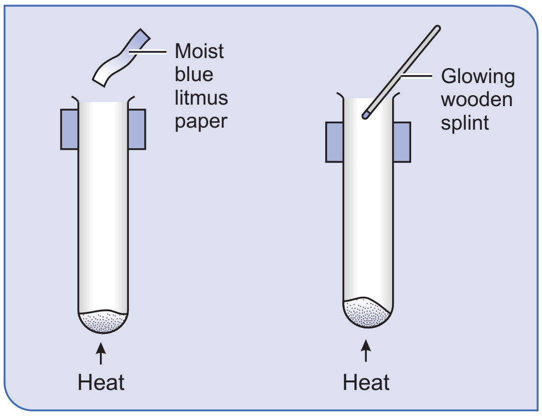
Activity	Observation	Inference
Take a few crystals of copper sulphate in a dry test tube. Heat the test tube over a flame for a few minutes.	The colour of copper sulphate changes from blue to white-water droplets start appearing on the upper inside portion of the test tube.	On heating copper sulphate loses water of crystallisation which appears as droplets. $CuSO_4.5H_2O \rightarrow CuSO_4 + 5H_2O$

Activity	Observation	Inference
Take a small amount of washing soda in a dry test tube. Heat the test tube over a flame for a few minutes.	When heated, it loses water of crystallisation and forms a white powdery mass.	White powdery sodium carbonate monohydrate is formed. $Na_2CO_3.10H_2O \rightarrow Na_2CO_3.H_2O + 9H_2O$

(c) Zinc nitrate, Copper nitrate and Lead nitrate

Experiment :

1. About two spatulas full of copper(II) nitrate, zinc nitrate and lead nitrate are added in three test tubes.
2. The colour of the nitrate salt is noted.
3. The nitrate salt is then heated strongly as shown in Figure.



Observations :

Nitrate salt	Colour of salt before heating	Colour of residue		Tests for gases evolved			Gases evolved
		Hot	Cold	Colour of gas	Glowing splint	Blue litmus paper	
Zinc nitrate	White	Yellow	White	Brown gas and colourless gas	Rekindles	Turns red	Oxygen and nitrogen dioxide
Copper nitrate	Blue	Black	Black	Brown gas and colourless gas	Rekindles	Turns red	Oxygen and nitrogen dioxide
Lead nitrate	White	Brown	Yellow	Brown gas and colourless gas	Rekindles	Turns red	Oxygen and nitrogen dioxide

Reactions :

$$\underset{\text{Colourless}}{2\,Zn(NO_3)_2\,(s)} \longrightarrow \underset{\left[\begin{array}{c}\text{Yellow when hot}\\ \text{white when cold}\end{array}\right]}{2\,ZnO(s)} + 4\,NO_2(g) + O_2(g)$$

$$\underset{\text{Colourless}}{2Pb(NO_3)_2(s)} \longrightarrow \underset{\text{Yellow}}{2PbO(s)} + 4NO_2(g) + O_2(g)$$

$$\underset{\text{Blue}}{2Cu(NO_3)_2(s)} \longrightarrow \underset{\text{Black}}{2CuO(s)} + 4NO_2(g) + O_2(g)$$

(d) Ammonium Chloride, Iodine, Ammonium dichromate

Ammonium chloride : Initial heating of ammonium chloride causes the salt to sublime. Further heating causes ammonium chloride to undergo thermal dissociation, to form ammonia and hydrogen chloride. Ammonia and hydrogen chloride recombine at the cooler upper parts of the test tube, and settle down as a white layer.

$$NH_4Cl(s) \underset{\text{when cooled}}{\overset{\text{when heated}}{\rightleftharpoons}} NH_3(g) + HCl\ (g)$$

Ammonia turns red litmus paper blue and hydrogen chloride changes blue litmus again to red.

Iodine : Iodine is a greyish brown crystalline solid, when heated it undergoes sublimation and changes to violet iodine vapours.

$$\text{Iodine crystal} \rightleftharpoons \text{Iodine vapour}$$

Ammonium dichromate : It is an orange coloured crystalline substance. When heated, it starts decomposing with sparks and evolution of heat. The reaction then proceeds with its own heat, and a greenish fluffy powder called chromic oxide is thrown out of the test tube.

The other products are water and nitrogen.

$$\underset{\substack{\text{Ammonium}\\\text{dichromate}\\\text{Orange}}}{(NH_4)_2Cr_2O_7(s)} \longrightarrow \underset{\substack{\text{Chromic}\\\text{oxide}\\\text{green}}}{Cr_2O_3(s)} + 4H_2O(vap) + N_2(g)$$

Experiment No. 2

Action of dilute sulphuric acid on following substances :

(1) a metal (2) a carbonate (3) a sulphide (4) a sulphite

Activity :

S. No.	Compound + dil. sulphuric acid	Observation	Test
1.	Metal + dil. sulphuric acid	Effervescence, colourless and odourless gas is evolved.	Apply a lit splint or ignite the gas evolved. A pop sound confirms hydrogen gas.
2.	Carbonate + dil. sulphuric acid	Brisk effervescence and a colourless odourless gas is evolved.	Gas can be tested by bubbling the gas produced through limewater, a dilute solution of calcium hydroxide, the limewater turns milky which confirms the presence of carbon dioxide gas.
3.	Sulphide + dil. sulphuric acid	Effervescence and a colourless gas with rotten eggs smell is evolved.	Gas evolved is tested with damp lead(II) ethanoate paper lead acetate, Rotten egg smell and the gas turns lead(II) ethanoate paper black which confirms H_2S gas.

S. No.	Compound + dil. sulphuric acid	Observation	Test
4.	Sulphite + dil. sulphuric acid	Colourless gas with smell of burning sulphur is evolved.	Gas evolved is passed through a concentrated, acidified solution of potassium dichromate ($K_2Cr_2O_7$). Solution changes from orange to green confirms the presence of sulphur dioxide gas.

Experiment No. 3

Apply the flame test to identify the metal in the unknown substance.
(a) a sodium salt. (b) a potassium salt (c) a calcium compound
Every metal gives a characteristic colour to the flame which helps in identification of the metal.

Procedure :

1. Make a paste of compound in conc. HCl
2. Now, hold the paste in the non-luminous zone of the flame or burner by means of a platinum wire loop.
3. Note the colour of the flame.

S. No.	Colour of the flame	Inference
1	Brick red or dull red	Calcium
2	Golden yellow	Sodium
3	Violet (Lilac)	Potassium
4	Blue green	Copper

Experiment No. 4

Simple experiments based on hard water and soft water, identification of hardness - simple softening-by heating temporary hard water, using washing soda and advantages of detergents over soap in hard water.

Water that produces lather with soap readily is called soft water. The water that does not produce lather with soap solution readily is called hard water. The hardness of water is due to the presence of bicarbonates, chlorides and sulphates of calcium and magnesium.

Activity 1

To study the action of soap with soft and hard water.

1. Take about 10 mL of distilled water (or rain water) and 10 mL of hard water (from a tube well or hand-pump) in separate test tubes.
2. Add a couple of drops of soap solution to both.
3. Shake the test tubes vigorously for an equal period of time and observe the amount of foam formed.

Observation :

In test tube containing distilled water, soap gives foam easily whereas in test tube having hard water, a curdy white ppt. is formed.

Conclusion :

Soap produce foam with soft water and curdy white ppt. with hard water.

Activity 2

To show that hard water produces foam with detergent but only curdy white ppt. with soap.

1. Take two test tubes with about 10 mL of hard water in each.

2. Add five drops of soap solution to one and five drops of detergent solution to the other.
3. Shake both test tubes for the same period.

Observation :

The test tube having soap solution a curdy white ppt. is formed but in test tube containing detergent produces a large amount of foam.

Conclusion :

Detergents can be used in both soft and hard water but soaps only in soft water.

Hardness of water is of two types:

(a) Temporary hardness : It is due to presence of bicarbonates of calcium and magnesium. It can be removed by boiling.

(b) Permanent hardness : It is due to presence of chlorides and sulphates of calcium and magnesium. It can be removed by adding washing soda.

Activity 3

Boil the given sample of water.

Temporary hardness in water can be easily removed by boiling. During boiling, the soluble $Mg(HCO_3)_2$ is converted into insoluble $Mg(OH)_2$ and $Ca(HCO_3)_2$ is changed to insoluble $CaCO_3$. It is because of high solubility product of $Mg(OH)_2$ as compared to that of $MgCO_3$, that $Mg(OH)_2$ is precipitated. These precipitates can be removed by filtration. Filtrate thus obtained will be soft water.

$$Mg(HCO_3)_2 \xrightarrow{\text{Heating}} Mg(OH)_2 + 2CO_2 \uparrow$$

$$Ca(HCO_3)_2 \xrightarrow{\text{Heating}} CaCO_3 \downarrow + H_2O + CO_2 \uparrow$$

Removal of hardness of water by adding washing soda :

Permanent harness of water can be removed by adding washing soda.

Activity 4

(a) Take two beakers and label them as A and B.

(b) Add washing soda to both the beakers.

(c) Filter the precipitate formed and add soap solution to both the beakers.

Observation :

Lather is formed with soap solution easily by both.

Conclusion :

Both temporary and permanent hardness of water is removed by adding washing soda.

The bicarbonates responsible for temporary hardness are converted to insoluble carbonates which can be removed by filtration.

$$Mg(HCO_3)_2 \xrightarrow{\text{Heating}} Mg(OH)_2 \downarrow + 2CO_2 \uparrow$$

$$Ca(HCO_3)_2 \xrightarrow{\text{Heating}} CaCO_3 \downarrow + H_2O + CO_2 \uparrow$$

Washing soda reacts with soluble calcium and magnesium chlorides and sulphates in hard water to form insoluble carbonates which can be removed by filtration.

$$MgSO_4 + Na_2CO_3 \rightarrow MgCO_3 \downarrow + Na_2SO_4$$

$$MgCl_2 + Na_2CO_3 \rightarrow MgCO_3 \downarrow + 2NaCl$$

$$CaSO_4 + Na_2CO_3 \rightarrow CaCO_3 \downarrow + Na_2SO_4$$

$$CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 \downarrow + 2NaCl$$

Advantages of detergents over soaps in hard water : Soaps cannot be used in hard water since calcium, magnesium ions present in hard water produce curdy precipitates of calcium and magnesium salts.

$$\underset{\text{soluble sodium stearate}}{2C_{17}H_{35}COONa} + \underset{\text{Calcium ions}}{Ca^{2+}} \rightarrow \underset{\text{Insoluble calcium stearate}}{(C_{17}H_{35}COO)Ca} + 2Na^+$$

$$\underset{\text{Sodium stearate}}{2C_{17}H_{35}COONa} + \underset{\text{Magnesium ions}}{Mg^{2+}} \rightarrow \underset{\text{Magnesium stearate}}{(C_{17}H_{35}COO)Mg} + 2Na^+$$

These insoluble soaps separate as scum in water and causes hindrance to wash because the precipitate adheres onto the fibre of the cloth as gummy mass. Thus, a lot of soap is wasted if water, is hard. On the other hand, Synthetic detergents can be used even in case of hard water.

Experiment No. 5

Find out the sources of pollution of water bodies in the locality. Suggest preventive measures to control it.

Carry out a survey for sources of water pollution. Sources of water pollution, particularly ground water pollution are group under two categories based on the origin of the pollutant.

(a) Point Source Pollution : Water pollution caused by point sources refers to the contaminants that enter the water body from a single, identifiable source like pipe or ditch. Point source pollutants can be discharges from sewage treatment plant, factories or a city storm drain.

(b) Non-point Source Pollution: Pollution caused by non-point sources refers to the contamination that does not originate from a single source. Non-point source pollution is the cumulative effect of small contaminants gathered in large area. Leaching of nitrogen compounds from agricultural land, storm water runoff over an agricultural land or a forest are examples of non-point source pollution.

Main water pollutants are sewage, metals such as Lead, Cadmium, Arsenic, Plastics, detergents etc.

On collecting samples of water, sources of water pollution can be categorized as :

(a) Municipal sources
(b) Industrial sources
(c) Agricultural sources
(d) Sources of underground pollution.

Preventive measures

Water pollution can be reduced with some efforts, such as

1. Identify industrial units that are the biggest polluters of river water. If NEQS regarding wastewater were strictly enforced, these industries would have to reduce and treat their waste water before disposal.
2. A regular qualitative and quantitative monitoring of fresh water resources.
3. Construct proper sanitary landfill sites.
4. Investigate ground water quality.
5. Provide government help for waste management by industries.
6. Throw refuse into garbage cans.
7. To prevent petroleum and other oily products to enter the sea.
8. To prevent the entry of radioactive substance in sea and ban on radioactive explosions in sea.
9. Avoid bathing, washing or other activities near the water source such as ponds, lakes etc.
10. People should be made aware of harms caused due to water pollution through radio, television and other media.

SUMMARY

- If a gas is colourless and odourless, neutral to litmus, gives squeaky pop sound with lit split is hydrogen.
- If a gas is colourless and odourless, neutral to litmus and reignites the flame when a glowing split is brought near it, it shows the gas present is oxygen gas.
- If a gas is colourless and odourless, changes moist blue litmus red, turns limewater milky, it shows the gas present is carbon dioxide gas.
- Greenish yellow colour gas with pungent odour, turns blue litmus red which later turns colourless, gives white precipitate with silver nitrate is chlorine gas.
- If a gas is colourless and pungent smell, changes moist blue litmus red but does not bleach it, gives white precipitate with silver nitrate shows that the gas present is hydrogen chloride.
- If a gas is colourless, smells like burning sulphur, changes moist blue litmus red and then bleaches it. Changes solution of potassium dichromate from orange to green shows that the gas present is sulphur dioxide.
- If the gas is colourless smells like rotten egg, changes moist blue litmus red, turns lead(II) ethanoate paper black shows the gas present is hydrogen sulphide.
- If a gas is colourless, has a pungent smell,turns moist red litmus blue, gives white clouds with HCl fumes, then it shows the presence of Ammonia gas.
- Action of heat on compounds :

Compound	Colour of metal carbonate before heating	Colour of metal carbonate after heating
Copper carbonate	Bluish green	White
Zinc carbonate	White	Yellow
Zinc nitrate	White	Yellow
Copper nitrate	Blue	Black

Lead nitrate	White	Brown yellow
Iodine	Greyish brown	Violet
Ammonium dichromate	Orange	Green

- **Action of dilute sulphuric acid:**

S. No.	Compound	Observation
1	Metal	Hydrogen gas is evolved.
2	Carbonate	Carbon dioxide gas is evolved.
3	Sulphide	Hydrogen sulphide gas
4	Sulphite	Sulphur dioxide gas is evolved.

- **Flame Test :**

S. No.	Colour of the flame	Inference
1	Brick red or dull red	Calcium
2	Golden yellow	Sodium
3	Violet (Lilac)	Potassium
4	Blue green	Copper

- Water that produces lather with soap readily is called soft water. The water that does not produce lather with soap solution readily is called hard water. The hardness of water is due to the presence of bicarbonates, chlorides and sulphates of calcium and magnesium.
- Hardness may be temporary and permanent. Temporary hardness can be removed by hardness. Permanent hardness can be removed by adding washing soda.

EXERCISE

SECTION A

1. **Identify the gas :**
   (a) Colourless and odourless gas, changes moist blue litmus red, turns limewater milky.
   (b) Colourless gas smells like burning sulphur, changes moist blue litmus red and then bleaches it.
   (c) Colourless, pungent smelling gas, turns moist red litmus blue, gives white clouds with HCl fumes shows the presence of ammonia gas.
2. **Match the element with its flame colour :**

S. No.	Colour of the flame	Inference
1	Brick red or dull red	Sodium
2	Golden yellow	Copper
3	Violet (Lilac)	Calcium
4	Blue Green	Potassium

3. How permanent hardness of water is removed by washing soda?
4. **Distinguish between with respect to action of heat on them :**
   (a) Copper carbonate and Zinc carbonate.
   (b) Lead nitrate and Copper nitrate
   (c) Iodine and Ammonium chloride.
5. **Which gas is evolved when :**
   (i) Metal react with dilute sulphuric acid.
   (ii) A sulphide with dilute sulphuric acid.

SECTION B

1. **Write the balanced chemical equation for :**
   (i) Lead nitrate is heated.
   (ii) Zinc carbonate is heated.
   (iii) Washing soda is heated.
2. How will you distinguish between chlorine and hydrogen chloride gas using litmus paper.
3. Write the preventive measures for control of water pollution.

4. **Complete the following table :**

Nitrate salt	Colour of salt before heating	Colour of residue	
		Hot	Cold
Zinc nitrate	White		White
Copper nitrate		Black	
Lead nitrate	White		

5. **Give a test for the identification of :**
   (i) Oxygen gas
   (ii) Ammonia gas
   (iii) Nitrogen dioxides gas.

Research Questions :

(a) Carry out a survey for sources of water pollution. Give at least three parameters to check the quality of water.

(b) Give an activity to distinguish between hard water and soft water.

(c) Classify the sources of water as point sources and non-point sources. Give examples. Also write preventive measures to control water pollution.

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Model paper 1 (Unsolved)

Paper 2 (Chemistry)

(Two hours)

- *Answers to this Paper must be written on the paper provided separately.*
- *You will not be allowed to write during the first 15 minutes.*
- *This time is to be spent in reading the question paper.*
- *The time given at the head of this Paper is the time allowed for writing the answers.*
- *Section I is compulsory. Attempt any four questions from Section II.*
- *The intended marks for questions or parts of questions are given in brackets [ ].*

SECTION—I (40 MARKS)

Compulsory (Attempt all questions from this Section)

Question 1

(a) Fill in the Blanks :

(i) The isotope of hydrogen having mass number 3 is ………… (Protium, Deuterium, Tritium)

(ii) Among the halogen, ………is most reactive. (Chlorine, Bromine, Fluorine)

(iii) Group 17 elements are known as ……….. (Halogens, Chalcogens, Noble gases)

(iv) Carbon dioxide gas can trap ………… radiations. (Infrared, ultraviolet, Visible)

(v) Pure water is …… (Neutral, acidic,basic).

(b) Choose the correct answer :

1. Gas pressure is caused by :
(a) Gas molecules heating up
(b) Gas molecules reacting with other gas molecules
(c) Gas molecules hitting the walls of a container
(d) Gas molecules hitting other gas molecules

2. The symbol of sodium is :
(a) Sa (b) Sd
(c) Na (d) Au

3. Carbon tetrachloride molecule has :
(a) Two covalent bonds
(b) One covalent bond
(c) Three covalent bonds
(d) None of the above

4. Which group in periodic table has maximum number of elements ?
(a) Group 1 (b) Group 2
(c) Group 3 (d) Group 4

5. Which of the following is not a source of potable water ?
(a) Well (b) River
(c) Lake (d) Sea

(c) Identify the substances underlined, in each of the following cases :

(i) **Element** that gives blue-green colour with flame.

(ii) **Element** having highest electronegativity.

(iii) An **element** which is tetravalent.

(iv) A **solution** in which no more solvent can be dissolved.

(v) The **bond** found in oxygen molecule.

(d) Write a balanced chemical equation for each of the following :

(i) Reaction of Zinc and Copper sulphate solution.

(ii) Action of steam on carbon.

(iii) Action of Sulphuric acid on Magnesium.

(iv) Industrial preparation of Hydrogen

(v) Formation of ozone

(e) State one relevant observation for each of the following reactions :

(i) Action of acid rain on Marble monuments.

(ii) Decomposition of calcium carbonate.

(iii) Action of KOH on Aluminum.

(iv) Action of hot water on Magnesium.

(v) Reaction of silver nitrate and sodium chloride.

(f) (i) Draw the Atomic orbital structure of each of the following :

1. Methane 2. Sodium chloride
3. Carbontetrachloride

(ii) Define the following words :

(a) Efflorescence
(b) Deliquescence

(g) (i) Convert 50°C to Kelvin.
(ii) Find the percent composition of each element in water.
(iii) What is the relative mass formula of sodium chloride? (Relative atomic masses : Na = 23, Cl = 35.5)

(h) The elements of the second group of the periodic table are Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba), and Radium (Ra).
(a) What is the group called ?
(b) Which element has the highest atomic size ?
(c) Which is the least metallic element ?
(d) Which element has the highest ionisation energy ?
(e) What is the nature of the elements: metallic or non-metallic ?

SECTION—II (40 MARKS)

*(Attempt any **four** questions from this Section)*

Question 2

(a) Identify the type of Reactions :
(i) Reaction of magnesium and oxygen to form magnesium oxide.
(ii) Heating of calcium carbonate.
(iii) Reaction of Potassium hydroxide and nitric acid to form Potassium nitrate and water.
(iv) Reaction of Zinc oxide and carbon to form zinc and carbon monoxide

(b) How many covalent bonds are present in :
(i) Oxygen
(ii) Nitrogen

(c) Answer the following questions :
(i) Give two examples of hydrated compounds.
(ii) How will you distinguish between Copper Carbonate, Zinc Carbonate by heating ?

(d) Identify the gas from the following observation :
(i) Greenish yellow colour gas with pungent odour. It turns blue litmus red which later turns colourless and gives white precipitate with silver Nitrate.
(ii) Colourless gas smells like burning sulphur, changes moist blue litmus red and then bleaches it. Changes colour of acidified solution of potassium dichromate from orange to red.

Question 3 :

(a) Draw an electron dot diagram to show the formation of the following :
(i) Water
(ii) Sodium chloride

(b) (i) State Boyle's law.
(ii) Write relation between Celsius and Kelvin scale of temperature.

(c) Account for the following facts :
(i) Though lead is above hydrogen in the activity series, it does not reacts with dilute hydrochloric acid or dilute sulphuric acid.
(ii) Potassium and Sodium are not used to react with dilute hydrochloric acid or dilute sulphuric acid in the laboratory preparation of hydrogen.

Question 4 :

(a) Define and give an example :
(i) Temporary hardness
(ii) Permanent hardness
(iii) Desiccating Agent
(iv) Amorphous crystal
(v) Dehydrating agents

(b) Write equation for the following reaction :
(i) Chlorine is passed over heated iron.
(ii) When potassium chlorate is strongly heated.
(iii) When ammonia is passed over heated copper oxide.

(c) Explain any 3 uses of hydrogen and give the reasons for its use.

Question 5 :

(a) Write the formula of the following compounds :
(i) Magnesium chloride
(ii) Copper carbonate
(iii) Barium sulphate
(iv) Sodium hydroxide

(b) Which test would be appropriate in each of these cases :
(i) To find out whether a given solution is saturated or unsaturated.
(ii) To distinguish between red mercuric oxide and red lead.
(iii) To identify a gas as being chlorine.
(iv) To prove that air contains water vapour.

(c) Write two advantages and disadvantages of soft water and hard water.

Question 6 :

(a) With reference to the Modern Periodic Table, name the following :

(i) An alkaline earth metal found in the fourth period.

(ii) A halogen of third period

(iii) The number of elements in the fifth period.

(iv) The valence shell of the elements of the third period.

(b) (i) State Charles law. Write mathematical equation also.

(ii) Give an example of molecule having dipole-dipole interaction.

(c) Give a balanced chemical equation for each of the following :

(i) Tin(II) chloride is heated with concentrated nitric acid.

(ii) A reaction where water acts as a catalyst.

Question 7 :

(a) (i) Name two important process which generate or release carbon dioxide into the atmosphere.

(ii) Name two process which removes carbon dioxide from the atmosphere.

(b) Give reasons for the following :

(i) Water can dissolve a large amount of substance.

(ii) Tap water is healthier than rain water.

(iii) Although carbon dioxide is fairly soluble in water, it can still dissolve in water of soda bottles.

(c) Identify the substance :

(i) Compound responsible for ozone hole.

(ii) One of the major greenhouse gases.

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Model paper 2 (Unsolved)

Paper 2 (Chemistry)

(Two hours)

- *Answers to this Paper must be written on the paper provided separately.*
- *You will not be allowed to write during the first 15 minutes.*
- *This time is to be spent in reading the question paper.*
- *The time given at the head of this Paper is the time allowed for writing the answers.*
- *Section I is compulsory. Attempt any four questions from Section II.*
- *The intended marks for questions or parts of questions are given in brackets [ ].*

SECTION—I (40 MARKS)

Compulsory (Attempt all questions from this Section)

Question 1

(a) Fill in the Blanks :

(i) If the atomic number of an element is 12 then its atom contains 12......
(Electrons, protons, neutrons).

(ii) ....reaction is the reaction in which a more reactive metal displaces a less reactive metal.
(Addition, Displacement, Neutralisation)

(iii) Law of Octaves is given by ....
(Newlands, Dalton, Mendeleev)

(iv) .... gas is found in Stratosphere.
(Methane, Nitrogen, ozone)

(v) Adding more solute to a solvent .... its boiling point. (Increases, decreases, neutral)

(b) Choose the correct answer :

1. In Avogadro's law volume is directly proportional to :
(a) Pressure (b) Number of moles
(c) Volume (d) Temperature

2. For writing chemical formula of a compound, which of the following should be written first ?
(a) Cation (b) Anion
(c) Neutral atom (d) None

3. The number of valence electrons in Mg^{2+} ion is :
(a) 14 (b) 10
(c) 12 (d) 0

4. Which of the following pair of elements belongs to same group ?
(a) Cu, Ag, Au (b) Li, Be, Al
(c) Na, Mg, Al (d) O, S, Cl

5. During boiling for removal of Temporary hardness of water, Magnesium bicarbonate is changed into
(a) Magnesium oxide
(b) Magnesium sulphate
(c) Magnesium hydroxide
(d) Magnesium carbonate

(c) Identify the substances underlined, in each of the following cases :

(i) <u>Gas</u> which changes the colour of Potassium Dichromate from orange to green.

(ii) <u>**Atoms**</u> of some elements which have same atomic number but different number of neutrons or different atomic mass.

(iii) A mark or <u>**character**</u> used as a conventional representation of an object

(iv) <u>**Property**</u> of a substance to absorb moisture from the air.

(v) The <u>**energy**</u> of an object because of its motion.

(d) Write the balanced chemical equation for each of the following :

(i) Reaction of Ammonia and Hydrochloric acid.

(ii) Reaction of Sodium oxide and water.

(iii) Action of hydrochloric acid on Zinc.

(iv) Action of Sodium hydroxide on lead.

(v) Formation of Nitric acid from Nitrogen oxides.

(e) State one relevant observation for each of the following reactions :

(i) Action of heat on Copper carbonate.

(ii) Action of heat on washing soda.

(iii) Reaction of dilute sulphuric acid and a metal carbonate.

(iv) Action of steam on iron.

(v) Reaction of cold water with sodium.

(f) (i) Draw the Atomic orbital structure of each of the following :

1. Hydrogen
2. Hydrogen Chloride
3. Ammonia

(ii) Define the following words :

(a) Hydrated substances
(b) Water of Crystallization

(g) Solve each of these numericals :

(i) For the compound $Ca(OH)_2$ or calcium hydroxide, find the percent composition of the element Calcium (Ca)
(ii) What percentage of the mass of ammonium nitrate is nitrogen? (The formula for ammonium nitrate is NH_4NO_3, Relative atomic masses : H = 1, O = 16, N = 14).
(iii) Convert 45°C to Fahrenheit.

(h) The following questions refer to the periodic table :

(i) Name the last element of the period 4.
(ii) How many elements are in the third period ?
(iii) Name the element which has the zero electron affinity.
(iv) Name the element which has the highest electro-negativity.
(v) Name the element which is liquid.

SECTION—II (40 MARKS)

*(Attempt any **four** questions from this Section)*

Question 2

(a) Balance the following equations :

(a) $Ca(HCO_3)_2 + Ca(OH)_2 \rightarrow CaCO_3 + H_2O$
(b) $Cu + H_2SO_4 \rightarrow CuSO_4 + SO_2 + H_2O$
(c) $Mg + O_2 \rightarrow MgO$
(d) $Na_2O + H_2O \rightarrow NaOH$

(b) Give the symbol of :

(i) Gold
(ii) Magnesium

(c) Answer the following questions :

(i) Give a test to distinguish between Zinc nitrate and Lead nitrate.
(ii) Give the reaction for removal of permanent hardness of water by washing soda

(d) Identify the element from the following colour of flame test :

(i) Blue green colour
(ii) Golden yellow

Question 3 :

(a) Show with the help of diagram number of electrons, protons and neutrons in Deuterium and Tritium.

(b) (i) Which fuel can be used in vehicles to reduce pollution ?
(ii) Which gas contributes largely in global warming ?

(c) Give reason for the following statements :

(i) Silver nitrate solution is kept in coloured reagent bottles in the laboratory.
(iii) An element has atoms with differing mass numbers.

Question 4 :

(a) Define the following terms :

(i) Sublimation
(ii) Boyle's Law
(iii) Displacement reaction
(iv) Water of Crystallization

(b) Write equation for the following reaction :

(i) Conversion of cobalt to a cation.
(ii) Chlorine gas is bubbled through water.
(iii) Calcium hydroxide reacts with Ammonium chloride.

(c) Give two examples to show how hydrogen can be prepared by— (Give only the equations.)

(i) Using an alkali
(ii) From reaction between a metal and an acid
(iii) Bosch process

Question 5 :

(a) Balance the following equation :

(i) $CuO + NH_3 \rightarrow Cu + H_2O + N_2$
(ii) $Na + H_2O \rightarrow NaOH + H_2$
(iii) $N_2 + H_2 \rightarrow NH_3$
(iv) $Fe_2O_3 + C \rightarrow Fe + CO$

(b) State the observation :

(i) When burning magnesium is placed in the jar of oxygen.
(ii) When Sulphur is burnt.
(iii) When dry slaked lime is sprinkled in a jar of chlorine.
(iv) When carbon dioxide is passed through lime water

(c) A balloon takes up 625L at 0°C. If it is heated to 80°C, what will its new volume be ?

Question 6 :

(a) Give the position of the following elements in periodic table : (Group and Period)

(i) Chlorine (ii) Carbon
(iii) Potassium (iv) Oxygen
(v) Bromine

(b) (i) State Gay Lussac's Law.

(ii) The pressure in a sealed can of gas is 235 kPA when it sits at room temperature (20°). If the can is warmed to 48°C, what will the new pressure inside the can be ?

(c) (i) An element with electronic configuration 2.8.4.

(ii) An element which forms amphoteric oxide.

(iii) What is valency ?

Question 7 :

(a) Name the following :

(i) The change of state in which naphthalene changes into vapour.

(ii) A homogeneous mixture of a liquid and a solid.

(iii) One substance which is `deliquescent'.

(iv) The element which does not contain any neutron in its nucleus.

(b) Solve the following numercial problems related to Gas Laws :

(i) The volume of a certain gas was found 800 cm^3, when the pressure was 760 mm of mercury. If the pressure increases by 25%, find the new volume of the gas.

(ii) Sulphur dioxide occupies a volume of 512 cm^3 at S.T.P. Find its volume at 27°C and at a pressure of 720mm of mercury.

(c) (i) Name two acids responsible for acid rain.

(ii) Give two examples of monuments which are affected by Acid rain.

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APPENDIX

ELEMENTS, THEIR SYMBOL, ATOMIC NUMBER AND MOLAR MASS

Element	*Symbol*	*Atomic number*	*Molar mass/ (g mol^{-1})*	*Element*	*Symbol*	*Atomic number*	*Molar mass/ (g mol^{-1})*
Actinium	Ac	89	227.03	Hassium	Hs	108	(269)
Aluminium	Al	13	26.98	Helium	He	2	4.00
Americium	Am	95	(243)	Holmium	Ho	67	164.93
Antimony	Sb	51	121.75	Hydrogen	H	1	1.0079
Argon	Ar	18	39.95	Indium	In	49	114.82
Arsenic	As	33	74.92	Iodine	I	53	126.90
Astatine	At	85	210	Iridium	Ir	77	192.2
Barium	Ba	56	137.34	Iron	Fe	26	55.85
Berkelium	Bk	97	(247)	Krypton	Kr	36	83.80
Beryllium	Be	4	9.01	Lanthanum	La	57	138.91
Bismuth	Bi	83	208.98	Lawrencium	Lr	103	(262.1)
Bohrium	Bh	107	(264)	Lead	Pb	82	207.19
Boron	B	5	10.81	Lithium	Li	3	6.94
Bromine	Br	35	79.91	Lutetium	Lu	71	174.96
Cadmium	Cd	48	112.40	Magnesium	Mg	12	24.31
Caesium	Cs	55	132.91	Manganese	Mn	25	54.94
Calcium	Ca	20	40.08	Meitnerium	Mt	109	(268)
Californium	Cf	98	251.08	Mendelevium	Md	101	258.10
Carbon	C	6	12.01	Mercury	Hg	80	200.59
Cerium	Ce	58	140.12	Molybdenum	Mo	42	95.94
Chlorine	Cl	17	35.45	Neodymium	Nd	60	144.24
Chromium	Cr	24	52.00	Neon	Ne	10	20.18
Cobalt	Co	27	58.93	Neptunium	Np	93	(237.05)
Copper	Cu	29	63.54	Nickel	Ni	28	58.71
Curium	Cm	96	247.07	Niobium	Nb	41	92.91
Dubnium	Db	105	(263)	Nitrogen	N	7	14.0067
Dysprosium	Dy	66	162.50	Nobelium	No	102	(259)
Einsteinium	Es	99	(252)	Osmium	Os	76	190.2
Erbium	Er	68	167.26	Oxygen	O	8	16.00
Europium	Eu	63	151.96	Palladium	Pd	46	106.4
Fermium	Fm	100	(257.10)	Phosphorus	P	15	30.97
Fluorine	F	9	19.00	Platinum	Pt	78	195.09
Francium	Fr	87	(223)	Plutonium	Pu	94	(244)
Gadolinium	Gd	64	157.25	Polonium	Po	84	210
Gallium	Ga	31	69.72	Potassium	K	19	39.10
Germanium	Ge	32	72.61	Praseodymium	Pr	59	140.91
Gold	Au	79	196.97	Promethium	Pm	61	(145)
Hafnium	Hf	72	178.49	Protactinium	Pa	91	231.04

Element	*Symbol*	*Atomic number*	*Molar mass/ (g mol^{-1})*
Radium	Ra	88	(226)
Radon	Rn	86	(222)
Rhenium	Re	75	186.2
Rubidium	Rb	37	85.47
Ruthenium	Ru	44	101.07
Rutherfordium	Rf	104	(261)
Samarium	Sm	62	150.35
Scandium	Sc	21	44.96
Seaborgium	Sg	106	(266)
Selenium	Se	34	78.96
Silicon	Si	14	28.08
Silver	Ag	47	107.87
Sodium	Na	11	22.99
Strontium	Sr	38	87.62
Sulphur	S	16	32.06
Tantalum	Ta	73	180.95
Technetium	Tc	43	(98.91)
Terbium	Tb	65	158.92
Thallium	Tl	81	204.37
Thorium	Th	90	232.04
Thulium	Tm	69	168.93
Tin	Sn	50	118.69
Titanium	Ti	22	47.88
Tungsten	W	74	183.85
Uranium	U	92	238.03
Vanadium	V	23	50.94
Xenon	Xe	54	131.30
Ytterbium	Yb	70	173.04
Yttrium	Y	39	88.91
Zinc	Zn	30	65.37
Zirconium	Zr	40	91.22

The value given in parentheseis is the molar mass of the isotope of largest known half-life.

CHEMICAL NAME, FORMULA AND COMMON NAME OF COMPOUNDS

S.No.	CHEMICAL NAME	FORMULA	COMMON NAME
1	Acetic acid	CH_3COOH	Vinegar
2	Acetone	CH_3COCH_3	Acetone
3	Ethyne	C_2H_2	Acetylene gas
4	Alum	$K_2SO_4.Al_2(SO_4)_3.24H_2O$	Ammonium aluminium sulphate
5	Aluminium	Al	Aluminium
6	Aluminium hydroxide	$Al(OH)_3$	Alumina hydrate
7	Aluminium oxide	Al_2O_3	Alumina
8	Aluminium sulphate	$Al_2(SO_4)_3$	Flocculating Powder
9	Ammonia	$NH_3(aq)$;	Ammonia,
10	Ammonium hydroxide	NH_4OH	—
11	Ammonium bicarbonate	$(NH_4)HCO_3$	Salt of Hartshorn, Baker's Ammonia
12	Ammonium bromide	NH_4Br	—
13	Ammonium carbonate	$(NH_4)_2CO_3$	Smelling Salts, Salt of Hartshorn,
14	Ammonium chloride	NH_4Cl	Soldering Flux
15	Ammonium nitrate	NH_4NO_3	Ammonium nitrate
16	Ascorbic acid	$C_6H_8O_6$	Vitamin C
17	Sodium bicarbonate	$NaHCO_3$	Baking soda
18	Sodium hypochlorite or hydrogen peroxide	NaClO	Bleach (liquid)
19	Calcium hypochlorite	$Ca(ClO)_2$	Bleaching powder
20	Sodium tetraborate decahydrate	$Na_2B_4O_7{\cdot}10H_2O$	borax
21	Copper sulphate	$CuSO_4$	Blue vitriol
22	Stannic chloride	$SnCl_3$	Butter of tin
23	Calcium carbonate	$CaCO_3$	Limestone, Carbonate of Lime
24	Calcium chloride	$CaCl_2$	Laundry Aid/Road Salt/De-Icer
25	Calcium hydroxide	$Ca(OH)_2$	Slaked Lime, garden lime
26	Calcium hypochlorite	$Ca(ClO)_2$	—
27	Calcium nitrate	$Ca(NO_3)_2{\cdot}4H_2O$	Norwegian saltpeter
28	Calcium oxide	CaO	Lime, Quicklime
29	1,7,7-Trimethyl bicyclo [2.2.1] heptan-2-one	$C_{10}H_{16}O$	camphor
30	carbon	C	Soot, Graphite, Graphene, Carbon-nanotubes, Fullerenes, Diamond, Charcoal
31	Sodium hydroxide	NaOH	Caustic soda
32	Potassium hydroxide	KOH	Caustic potash
33	Carbolic acid	C_6H_6O	phenol
34	Sodium nitrate	$NaNO_3$	Chile saltpeter
35	Calcium sulphate [hemihydrates]	$2CaSO_4.H_2O$	Plaster of Paris
36	Carboxyl chloride	$COCl_2$	Phosgene
37	Carbontetrachloride	CCl_4	Pyrene
38	Carbon dioxide (solid)	CO_2	Dry ice

39	Ethanol/ethyl alcohol	C_2H_5OH	
40	Ethoxyethane	$(C_2H_5)_2O$	Ether
41	1, 2-ethanediol	CH_2OHCH_2OH	ethylene glycol
42	Formic acid	CH_2O_2	methanoic acid
43	Formaldehyde	$HCHO$	Formalin
44	Ferrous ammonium sulphate	$(NH_4)_2Fe(SO_4)_2{\cdot}6H_2O$	Mohrs salt
45	Fructose	$C_6H_{12}O_6$	Fruit Sugar
46	Glucose	$C_6H_{12}O_6$	
47	Ferrous sulphate	$FeSO_4$	Green vitriol
48	Calcium sulphate	$CaSO_4.2H_2O$	gypsum
49	Hydrochloric acid	HCl	
50	Hydrofluoric acid	HF	
51	Hypochlorous acid	$HClO$	Laundry Bleach
52	Hydrogen peroxide	H_2O_2	
53	2-Hydroxypropanoic acid	$C_3H_6O_3$	lactic acid
54	Magnesium hydroxide	$Mg(OH)_2$	Milk of Magnesia
55	Magnesium silicate	$Mg_3Si_4O_{10}(OH)_2$	Talc
56	Magnesium sulphate	$MgSO_47H_2O$	Epsom salt
57	Magnesium chloride	$MgCl_2$	Nigari or Lushui
58	Maganese dioxide	MnO_2	Pyrolusite
59	Methane	CH_4	Natural gas
60	Methanol	CH_3OH	Methyl Alcohol / wood's spirit
61	Methyl salicylate	$C_6H_4(OH)COOCH_3$	Oil of Wintergreen
62	Calcium carbonate	$CaCO_3$	Marble/limestone/chalk
63	Naphthalene	$C_{10}H_8$	moth balls
64	Potassium aluminium sulfate	$KAl(SO_4)_2{\cdot}12H_2O$	Alum, aluminium potassium sulphate
65	Potassium carbonate	K_2CO_3	Potash
66	Potassium bromide	KBr	—
67	Potassium chloride	KCl	Salt substitute
68	Potassium chromium sulphate	$KCr(SO_4)_2{\cdot}12H_2O$	potassium chrome alum
69	Potassium dichromate	$K_2Cr_2O_7$	—
70	Potassium nitrate	KNO_3	Saltpeter,
71	Potassium permanganate	$KMnO_4$	
72	Potassium sodium tartrate	$NaKC_4H_4O_6\ .4H_2O$	Rochelle salt
73	Silicon carbide	SiC	Carborundum
74	Silicon dioxide	SiO_2	Sand, silica
75	Silver nitrate	$AgNO_3$	Lunar caustic
76	Sodium carbonate	Na_2CO_3	Washing soda, soda ash
77	Sodium chloride	$NaCl$	Common Salt
78	Sucrose	$C_{12}H_{22}O_{11}$	Sugar
79	Sulphuric acid	H_2SO_4	
80	Tartaric acid	$C_4H_6O_6$	
81	Water	H_2O	—
82	Zinc sulphate	$ZnSO_4$	White vitriol

IUPAC, NOMENCLATURE SUMMARY

Group A : Functional Groups Indicated By Prefix Or Suffix

Family of Compound	Structure	Prefix	Suffix
Carboxylic acid	R—C(=O)—OH	carboxy-	-oic acid (carboxylic acid)
Aldehyde	R—C(=O)—H	oxo- (formyl)	-al (carbaldehyde)
Ketone	R—C(=O)—R	oxo-	-one
Alcohol	R—O—H	hydroxy-	-ol
Amine	R—N	amino-	-amine

Group B—Functional Groups Indicated By Suffix Only

Family of Compound		Prefix	Suffix
Alkane	—C—C—	—	-ane
Alkene	—C=C—	--------	-ene
Alkyne	—C≡C—	--------	-yne

Group C—Substituents Indicated by Prefix Only

Substituent		Prefix	Suffix
Alkyl	R—	alkyl-	----------
Alkoxy	R—O—	alkoxy-	----------
Halogen	F—	fluoro-	----------
	Cl—	chloro-	----------
	Br—	bromo-	----------
	I—	iodo-	----------

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GLOSSARY

- **Acid :** A compound that, when dissolved in water, gives a pH of less than 7.0 or a compound that donates a hydrogen ion.
- **Activation energy :** The minimum energy that must be input to a chemical system.
- **Addition reaction :** When two or more molecules combine to make a larger one.
- **Aeration :** The mixing of air into a liquid or solid.
- **Alkali metals :** The metals of Group 1 on the periodic table.
- **Alkaline earth metals :** The metals of Group 2 on the periodic table.
- **Allotropy :** Elements that can have different structures (and therefore different forms), such as Carbon (diamonds, graphite, and fullerene).
- **Anion :** Negatively charged ions.
- **Anode :** The positive side of a dry cell battery or a cell.
- **Atom :** A chemical element in its smallest form, and is made up of neutrons and protons within the nucleus and electrons within the nucleus.
- **Atomic number :** The number representing an element which corresponds with the number of protons within the nucleus.
- **Atomic orbital :** The region where the electron of the atom may be found.
- **Avogadro's number :** It is the number of particles in a mole of a substance.
- **Barometer :** a device used to measure the pressure in the atmosphere.
- **Base :** A substance that accepts a proton and has a high pH; a common example is sodium hydroxide (NaOH).
- **Boiling :** The phase transition of liquid vaporizing.
- **Boiling point :** The temperature at which the substance vapour pressure is equal to its atmospheric pressure.
- **Bond :** The attraction between atoms that enables the formation of a chemical compound.
- **Boyle's law :** For a given mass of gas at constant temperature, the volume varies inversely with the pressure.
- **Burette** (also buret) **:** Glassware used to dispense specific amounts of liquid when precision is necessary (e.g. titration).
- **Catalyst :** A chemical (element or compound) used to speed up a reaction, but is regenerated at the end of the reaction, without undergoing any change in itself.
- **Cation :** Positively charged ion.
- **Chemical reaction :** The change of one or more substances into another or multiple substances.
- **Colloid :** Mixture of evenly dispersed substances, such as milk.
- **Compound :** A substance that is made up of two or more chemically bonded elements.
- **Condensation :** The phase change from gas to liquid.
- **Conductor :** Material that allows electric flow more freely.
- **Covalent bond :** Chemical bond that involves sharing electrons.

- **Crystal :** A solid that is packed with ions, molecules or atoms in an orderly fashion.
- **Deionization :** The removal of ions, and in water's case mineral ions such as sodium, iron and calcium.
- **Density :** The amount of mass per unit volume. $d = m/V$.
- **Deposition :** Settling of particles within a solution or mixture.
- **Double bond :** Sharing of two pairs of electrons (in a covalent bond).
- **Diatomic :** Compound which consists of two atoms.
- **Electrolyte :** A solution that conducts a certain amount of current and can be split categorically as weak and strong electrolytes.
- **Electrochemical cell :** The cell where reaction between an anode and cathode takes place leading to an electro-chemical reaction.
- **Electron :** A subatomic particle with a net charge that is negative.
- **Electron shells :** An orbital around the atom's nucleus that has a fixed number electrons (usually two or eight).
- **Electric charge :** A measured property (Coulombs) that determines electromagnetic interaction.
- **Element :** An atom that is defined by its atomic number.
- **Energy :** Ability to do work.
- **Enzyme :** A protein that speeds up (catalyses) a reaction.
- **Freezing :** Phase transition from liquid to solid.
- **Gas :** Particles that fill their container though have no definite shape or volume.
- **Group :** The elements in a column of the periodic table. A family of elements.
- **Halogens :** Group 17 of the Periodic Table and are all non-metals.
- **Heat :** Energy transferred from one system to another by thermal interaction.
- **IUPAC :** International Union of Pure and Applied Chemistry.
- **Insulator :** Material that resists the flow of electric current.
- **Ion :** A molecule that has gained or lost one or more electrons.
- **Ionic bond :** Electrostatic attraction between oppositely charged ions.
- **Ionization :** The breaking up of a compound into separate ions.
- **Joule :** The SI unit of energy, defined as a newton-meter.
- **Kelvin :** A unit of measure for temperature based upon an absolute scale.
- **Lattice :** Unique arrangement of atoms or molecules in a crystalline liquid or solid.
- **Liquid :** A state of matter in which cohesive force is less than or equal to the separating force.
- **Metal :** Chemical element that is a good conductor of both electricity and heat and forms cations and ionic bonds with non-metals.
- **Melting :** The phase change from a solid to a liquid.
- **Metalloid :** A substance possessing both the properties of metals and non-metals.
- **Molecule :** A chemically bonded number of atoms that are electrically neutral.
- **Neutron :** A neutral unit or subatomic particle that has no net charge.

- **Nucleus :** The centre of an atom made up of neutrons and protons, with a net positive charge.
- **Noble gases :** Group 18 elements, whose outer electron shell is completely filled.
- **Non-metal :** An element which is not metallic.
- **pH :** The measure of acidity (or basicity) of a solution.
- **Plasma :** State of matter similar to gas in which a certain portion of the particles are ionized.
- **Proton :** A positive unit or subatomic particle that has a positive charge.
- **Radiation :** Energy in the form of waves or subatomic particles when there is a change from high energy to low energy states.
- **Salts :** Ionic compounds composed of anions and cations.
- **Semiconductor :** An electrically conductive solid that is between a conductor and an insulator.
- **Single bond :** Sharing of one pair of electrons.
- **Solute :** The part of the solution that is mixed into the solvent (NaCl in saline water).
- **Solution :** Homogeneous mixture made up of multiple substances. It is made up of solutes and solvents.
- **Solvent :** The part of the solution that dissolves the solute (H_2O in saline water).
- **Sublimation :** A process of solid transforming into gas and vice-versa without undergoing to liquid state.
- **Subatomic particles :** Particles that are smaller than an atom; examples are protons, neutrons and electrons.
- **Substance :** Material with definite chemical composition.
- **Triple bond :** The sharing of three pairs of electrons within a covalent bond (example N_2).
- **Valence electron :** The outermost electrons of an atom, which are located in electron shells.
- **Valence bond theory :** The theory which explains the chemical bonding within molecules by discussing valencies, the number of chemical bonds formed by an atom.
- **Valency :** The combining capacity of an element.
- **Van der Waals force :** One of the forces (attraction/repulsion) between molecules.
- **Vapour :** When a substance is below the critical temperature while in the gas phase.
- **Vapour pressure :** Pressure of vapour over a liquid at equilibrium.
- **Vaporization :** Phase change from liquid to gas.
- **Viscosity :** The resistance of a liquid to flow (oil).
- **Zuitter ion :** A molecule or ion having separate positively and negatively charged groups within the same molecule.

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