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CHEMISTRY

FOR CLASS – IX

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PREFACE

The century we have stepped in, is the century of Science and technology. The modern disciplines of Chemistry are strongly influencing not only all the branches of science but each and every aspect of human life.

To keep the students, abreast with the recent knowledge, it is must that the curricula, at all the levels, be updated regularly by introducing the rapid and multidirectional development taking place in all the branches of Chemistry.

The recent book of Chemistry for Class–IX has been written in this preview and in accordance with the revised curriculum prepared by Ministry of Education, Govt of Pakistan, Islamabad reviewed by independent team of Directorate of Curriculum Assessment and Research, Jamshoro sindh. Keeping in view the importance of Chemistry, the topics have been revised and re-written according to the need of the time.

This book consists of 8 chapters which have been thoroughly revised and re-written to meet the requirement of the curriculum. Special emphasis has also been paid to the applied aspect including the impact of Chemistry on daily life. Attention has also been focused on the different branches of Chemistry. Being a part of modern world, the aspects and problems of country are also discussed.

Among the new editions are the introductory paragraphs, information boxes, summaries and a variety of extensive exercises which I think will not only develop the interest but also add a lot to the utility of the book.

The Sind Textbook Board has taken great pains and incurred expenditure in publishing this book inspite of its limitations. A textbook is indeed not the last word and there is always room for improvement. While the authors have tried their level best to make the most suitable presentation, both in terms of concept and treatment, there may still have some deficiencies and omissions. Learned teachers and worthy students are, therefore, requested to be kind enough to point out the short comings of the text or diagrams and to communicate their suggestions and objections for the improvement of the next edition of this book.

In the end, I am thankful to Association for Academic Quality (AFAQ), our learned authors, editors and specialist of Board for their relentless service rendered for the cause of education.

Chairman
Sindh Textbook Board

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Teaching periods	=	12
Assessment period	=	3
Weightage	=	12

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1.3 Basic definitions
1.4 Chemical species
1.5 Chemical equation and balancing chemical equations
1.6 Mole and Avogadro's number
1.7 Chemical Calculations

STUDENTS LEARNING OUT COMES (SLO'S)

Students will be able to:

- Know the historical background of chemistry.
- Explain the contribution of Muslim scientists in the development of chemistry.
- Define chemistry and its importance in daily life.
- Identify and explain different branches of chemistry with the help of examples.
- Differentiate between main branches of chemistry.
- Distinguish between matter and a substance.
- Define ions, molecular ions, formula units and free radicals.
- Define atomic number, atomic mass and atomic mass unit.
- Differentiate among element, compound and mixture.
- Define relative atomic mass based on C-12.
- Differentiate empirical and molecular formula.
- Distinguish between atoms and ions.
- Differentiate between molecule and molecular ions.
- Distinguish between ions and free radicals.
- Classify the chemical species from the given examples.
- Relate gram atomic mass, gram molecular mass, and gram formula mass to mole.
- Describe that how Avogadro's number is related to mole of any substance.
- Identify the chemical equation in terms of moles.
- Calculation with balance equation using mole representative particles masses.
- Distinguish the terms gram atomic mass, gram molecular mass and gram formula mass
- Change atomic mass, molecular mass and formula into gram mass, gram molecular mass and gram formula mass.

INTRODUCTION

As we Know that word science comes from latin word "Scientia" which means knowledge, This knowledge is based on hypothesis observation and experiments of universal science. In this universal sciences chemistry purely deals with the matter which have mass and occupy space. Even from the table salt we use in cooking to electro chemical interaction of our human brain show the differences of substance because of the composition, structure, properties and interaction of matter.

The matter is undergoing changes continuously in nature as rusting of iron, evaporation of spirit and burning of coal are examples of reaction in which new substance are formed and energy is absorbed or released. All of these things are different due to the presence of different substances. Which are different by means of composition, properties, interaction , structure of matter.

The chemists use chemistry to explain occurrence and description of things. They investigate material, their interactions and propose theories to illuminate our understanding from a particle to galaxies.

1.1 HISTORICAL BACKGROUND OF CHEMISTRY:

Table 1.1 Time Chronology of Chemistry

Period /Timeline	Name of Scientists	Contribution/invention	Origin of scientist
384 - 322 B.C	Aristotle	Proposed idea of a substance as a combination of *matter* and *form*. Described theory of the Four Elements, i.e. fire, water, earth, air	Greek
347 - 428 B.C	Plato	Proposed term 'elements' as composition of organic and inorganic bodies with particular shape.	Greek
357 - 460 B.C	Democritus	Proposed the idea of atom, an indivisible particle of matter.	Greek
721 - 803 A.D	Jabir Ibne-Hayan	Invented experimental methods of nitric acid, hydrochloric acid and white lead. Extraction of metals from their ores and dyeing clothes.	Muslim

862-930 A.D	Al-Razi	Prepared ethyl alcohol by fermentation process.	Muslim
973-1048 A.D	Al-Beruni	Determined densities of different substances.	Muslim
980-1037 A.D	Ibne –Sina	Contributed in medicines, philosophy and astronomy.	Muslim
1627-1691 A.D	Robert Boyle	Put forward idea of chemistry as systematic investigation of nature. Discovered the gaseous law.	English
1728-1799 A.D	J. Black	Discovered carbon dioxide	Scottish
1733-1804 A.D	J.Priesly	Discovered oxygen,sulphur dioxide and hydrogen chloride.	English
1742-1786 A.D	Scheele	Discovered chlorine	German
1731-1810 A.D	Cavendish	Discovered hydrogen	British
1743-1794 A.D	Lavoisier	Discovered that oxygen is one fifth of air	French
1766-1844 A.D	John Dalton	Proposed atomic theory of matter	English
1778-1850 A.D	Gay-Lussac	Discovered that water is composed of two parts hydrogen and one part oxygen by volume. Discovered several chemical and physical properties of air and other gases,	French
1776-1856 A.D	Avogadro	Proposed Avogadro's lawthat equal volumes of gases under constant temperature and pressure contain equal number of molecules.	Italian

1746-1823 A.D	Jacques Charles	Described the gaseous law.	French
1741-1820 A.D	Petit	Determined the classical expression for the molar specific [heat capacity] of certain chemical elements.	French
1779-1848 A.D	J.J.Berzellius	Introduced symbols, formula and chemical equation to make study more systematic	Swedish
1824-1907 A.D	Mendeleve	Discovered periodic arrangement of elements.	Russian
1859-1927 A.D	Arrhenius	Proposed acid base theoryand ions dissociation.	Swedish
1791-1867 A.D	M.Faraday	Contributed to the study of electromagnetism and electrochemistry.	British
1856-1940 A.D	J.J.Thomson	Discovered the electron by experiments.	British
1885-1962 A.D	Neil Bohr	Proposed a theory for the hydrogen atom based on quantum theory	British
1871-1937 A.D	Rutherford	Postulated the nuclear structure of the atom. Discovered alpha and beta rays, and proposed the laws of radioactive decay.	Scottish
1887 - 1961 A.D	Schrodinger	Proposed Quantum mechanical model of atom	Australian

1892 - 1987 A.D	De Broglie	Proposed hypothesis about wave particle duality nature of electron.	French
1894 - 1974 A.D	Stendra Nath Bose	Proposed fourth state of matter	Indian
1879 - 1955 A.D	Elbert Eienstein	Proposed fourth state of matter	German
1961 - Alive	Eric Cornell	Synthesized the first Bose Einstein Condensate.	American
1951 - Alive	Carl weiman	Produced first bose Einstein Condensate	American

1.1.1 Definition of Chemistry:

Chemistry is the branch of science which deals with the properties, composition and structure of matter. Chemistry also deals with the changes involved in the matter .

1.1.2 Importance of Chemistry in daily life

Our planet earth has only life in the all planet of universe, due to existence of water (H_2O). The water is basic need of human, animals and plants. The chemical reactions take place in human, animals and plants. Disorder in these reactions may cause different diseases. Which may be over come with the help of chemistry.
The role of chemistry in daily life is unavoidable fact.

- Cooking, eating and digestion of food are purely chamical processes.
- Construction, cleaning and washing of our homes are dependable on chemistry.
- The production of fertilizers, glass, plastic synthetic fiber, polymer, ceramics, petroleum products, soaps, and detergents are based on chemistry.
- The diseases transmitted through impure drinking water as cholera, typhoid, dysentery, skin and eye infections can be controlled with the help of chlorine treatment to kill the pathogenic organism to obtain pure water.
- The chlorine is most important chemical which is used commercially to produce more than one thousand compounds which are used in chemical industry as bleaching agent, disinfectants, solvents, pesticides, refrigerates, PVC and drugs all are miracles of chemistry.

Test Yourself

- Identify and list down the chemistry related products in your home?
- How you can relate living things with chemistry?

Do You Know?

1.2 BRANCHES OF CHEMISTRY

As we know that chemistry is serving humanity everywhere in our environment. Due to its wide scope Chemistry is divided into following main branches:

1.2.1 Physical Chemistry

Physical chemistry is the branch of chemistry which deals with relationship between composition and physical properties of matter with the changes in them. It also deals with the laws and principles governing the combination of atoms and molecules in chemical reactions.

1.2.2 Organic Chemistry

Organic chemistry is the branch of chemistry which deals with hydrocarbons and their derivatives. Organic chemistry is the study of structure, properties, composition, reactions, and preparation of carbon-containing compounds, which include hydrocarbons except oxides, carbonates, bicarbonates and cynites. The gasoline, plastics, detergents, dyes, food additives, natural gas, and medicines are studied in the organic chemistry.

1.2.3 Inorganic Chemistry

Inorganic chemistry is the branch of Chemistry which deals with the study of all elements and their compounds except hydrocarbons. These compounds are generally obtained from nonliving things. It is applicable in all areas of chemical industry. Such as glass, cement, ceramics and metallurgy.

1.2.4 Biochemistry

Biochemistry is the branch of Chemistry which deals with the compounds of living organisms (plants and animals) their metabolism and synthesis in the living body such as carbohydrates, proteins and fats. Biochemistry helps us to understand how living things obtain energy from food. It tells that how disorder or deficiency of these biomolecules causes diseases. This branch is useful in medicine, agriculture and food science.

1.2.5 Industrial Chemistry

The branch of Chemistry which deals with the study of chemical processes involved in the chemical industries for the manufacture of synthetic products like fertilizers, glass, cement and medicines is called as industrial chemistry.

1.2.6 Nuclear Chemistry

Nuclear chemistry is the branch of Chemistry which deals with the radioactivity, nuclear processes and properties of radioactive substances. Radioactive elements are widely used in medicine as diagnostic tools and as a means of treatment, especially for cancer, preservation of food and generation of electric power through nuclear power reactors.

1.2.7 Environmental Chemistry

It is the branch of Chemistry which deals with the study of the interaction of chemical materials and their effect on the environment of animals and plants. Personal hygiene, pollution, health hazards are the important areas of environmental chemistry.

1.2.8 Analytical Chemistry

Analytical chemistry is the branch of chemistry which deals with separation and analysis of kind, quality and quantity of various components in given substance. It used in chromatography, electrophoresis and spectroscopy.

1.2.9 Medicinal Chemistry

The branch of Chemistry which deals with synthetic organic chemistry, pharmacology and various biological specialties. The medicinal chemistry is used in synthesis of chemicals, bioactive molecules (Drugs) and pharmaceutical agents.

1.2.10 Quantum Chemistry

The branch of Chemistry which deals with application, mechanics and experiments of physical models in chemical system. It is also called molecular quantum mechanics.

1.2.11 Green Chemistry

The branch of chemistry which deals with study of processes and designing products, which are composed of less hazardous substances. It is also known as sustainable chemistry.

Safer chemical (polyphenylsulfon), less hazardous chemical (poly carbons) and safer solvents are examples of green chemistry. The main purpose of this branch is to use waste material efficiently and improvement of energy efficiency in chemical industry.

Test Yourself

- In which branch of chemistry analysis of quality and quantity of compounds studied?
- What happens due to deficiency of biomolecules?
- Identify and list down examples of green chemistry in our environment?
- Differentiate among Medicinal chemistry and Biochemistry?

1.3 BASIC DEFINITIONS

1.3.1 Matter

Matter is all around us i.e. The air we are breathing, book we are reading and the stuff we touch and see. Matter is simply defined as anything that has mass and occupies space. It is found in three common states solid, liquid and gas. The plasma is also considered as fourth state of matter. The different states of matter are due to difference of energy in increasing order.

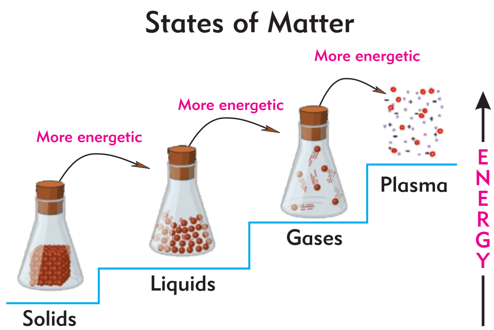
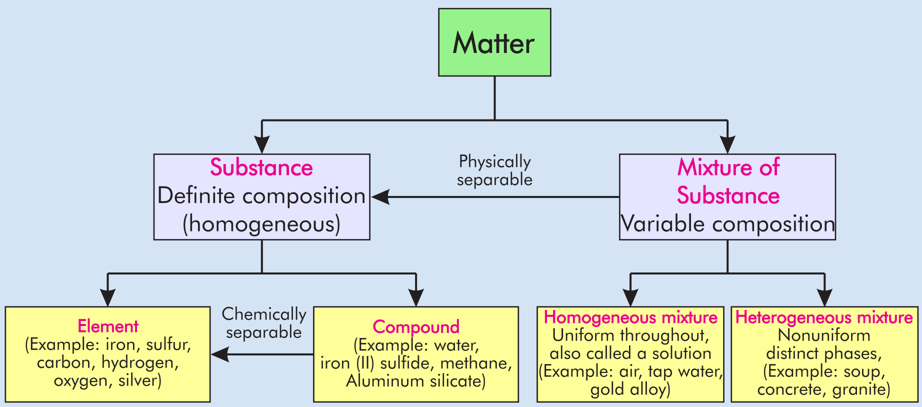


Fig 1.1 States of matter from solid to liquid, liquid to gases and gases to plasma due to increase in energy levels of the particles.

Do You Know?



1.3.2 Atom

Matter is made up of smallest particles which are known as atom. Atoms are the basic units of matter and define structure of elements. Now it is discovered that atoms are made up of three particles: protons, neutrons and electrons. Atoms are composed of even smaller particle as shown in figure 1.2. Where neutron and proton are situated in nucleus and electrons are revolving around the nucleus

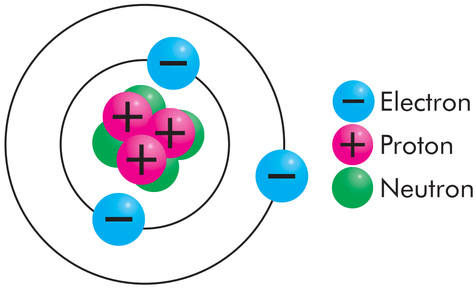
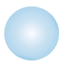
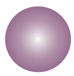
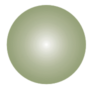
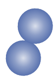
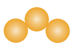
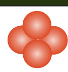
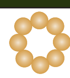


Fig 1.2 Particles of atom

1.3.3 Molecules

A molecule is the smallest particle in a chemical element or Compound that has the chemical properties of that element or Compound. Molecules are made up of atoms that are held together by chemical bonds. These bonds form as a result of the Sharing or exchange of electrons among atoms. Molecules are Mono, di and poly atomic molecules. Examples of mono,di and Poly atomic molecules are as follows.

Table 1.2 Examples of mono, di and poly atomic molecules

Monatomic elements					
					
Name	helium	argon	krypton	xenon	radon
Symbol	He	Ar	Kr	Xe	Rn
Diatomic molecules					
					
Name	nitrogen	oxygen	chlorine	bromine	iodine
Molecular formula	N_2	O_2	Cl_2	Br_2	I_2
Polyatomic molecules					
					
Name	ozone		phosphorus		sulphur
Molecular formula	O_3		P_4		S_8

1.3.4 Substance

A piece of matter in pure form is termed as a substance. Every substance has a fixed composition and specific properties. Every substance has physical and chemical properties. Examples of pure substances include tin, sulfur, diamond, water, pure sugar (sucrose), table salt (sodium chloride) and baking soda (sodium bicarbonate). Substances are elements and compounds.

Salt Diamond Sulphur

Fig 1.3 Examples of pure substance

1.3.5 Element

An Element is a substance made up of same type of atoms, Having same atomic number and cannot be decomposed into Simple substances by ordinary chemical reaction. Elements occur in nature in free or combined form in solid, Liquid and gaseous states. Till now 118 elements have been discovered. Majority of elements are solids as copper, gold, zinc etc. Very few elements are liquid as mercury and bromine. Few elements are gases as hydrogen, oxygen nitrogen etc. Elements are divided in to metals, nonmetals and metalloids on the basis of their properties.

Do you know?

Metal: solid material which is typically hard, shiny, malleable, fusible, and ductile, with good electrical and thermal conductivity (e.g. iron, gold, silver, aluminium, and alloys such as steel).

Nonmetal: is an element that doesn't have the characteristics of metal including: ability to conduct heat or electricity, luster, or flexibility. An example of a nonmetal element is carbon.

Metalloid: an element (e.g. arsenic, antimony, or tin) whose properties are intermediate between those of metals and non-metals i.e. semi conductor.

1.3.6 How to write Symbol?

Symbol is an abbreviation to represent the name of elements. A symbol is taken from the name of elements from English , Latin, Greek and German languages.

- Symbols are usually one or two letter long.
- Every symbol starts with capital letter as carbon with C or sulphur as S.
- If symbol is two letter then start with capital and second will be in small letter as He for helium, Cr for chromium.

Fig 1.4 Symbols of elements

The symbols of 30 elements in English derived from Latin, Greek and German are given in table 1.3

Table 1.3 Symbols of first 30 Elements

S. No	Names of Elements in English	Derived from Latin and Greek	Symbol
01	Hydrogen	Greek (root genes)	H
02	Hellim	Greek (Helios)	He
03	Lithium	Greek (lithos)	Li
04	Beryllium	Greek (beryllos)	Be
05	Boron	Latin (Busaq)	B
06	Carbon	Latin (Carbone)	C
07	Nitrogen	Greek (nitrumgenes)	N
08	Oxygen	Greek (oxygeinomes)	O
09	Flourine	Latin (fluor)	F
10	Neon	Greek (neos)	Ne
11	Sodium	Latin (Natrium)	Na
12	Magnesium	Greek (magnesium)	Mg
13	Aluminium	Latin (alumen)	Al
14	Silicon	Latin (silen)	Si
15	Phosphorous	Greek (Phoros)	P
16	Sulphur	Latin (sulohur)	S
17	Chlorine	Greek (Chloros)	Cl
18	Argon	Greek (argon)	Ar
19	Potassium	Latin (Kalium)	K
20	Calcium	Greek (Claix)	Ca
21	Scandium	Latin (scandia)	Sc
22	Titanium	Greek (titan)	Ti
23	Vanidium	Greek (vanadis)	V
24	Chromium	Greek (Chroma)	Cr
25	Mangnese	Greek (Magnesia)	Mn
26	Iron	Latin (Ferrum)	Fe
27	Cobalt	German (Kobold)	Co
28	Nichel	German (kupanickel)	Ni
29	Copper	Latin (Cuprum)	Cu
30	Zinc	German (zink)	Zn

1.3.7 What is valency?

The Combining power of an element with other element is called valency. The valency depends upon the number of electrons in the outermost shell. Valency is the number of electrons of an atom of an element can gain, lose or share. Some elements with their symbol and common valencies are given below in table 1.4

S. No	Elements	Symbol	Atomic Number	Valency
1	Hydrogen	H	1	1
2	Helium	He	2	0
3	Lithium	Li	3	1
4	Beryllium	Be	4	2
5	Boron	B	5	3
6	Carbon	C	6	4 2
7	Nitrogen	N	7	3
8	Oxygen	O	8	2
9	Flourine	F	9	1
10	Neon	Ne	10	0
11	Sodium	Na	11	1
12	Magnesium	Mg	12	2
13	Aluminium	Al	13	3
14	Silicon	Si	14	4
15	Phosphorus	P	15	3
16	Sulphur	S	16	2
17	Chlorine	Cl	17	1
18	Argon	Ar	18	0
19	Potassium	K	19	1
20	Calcium	Ca	20	2
21	Scanidium	Sc	21	3
22	Titanium	Ti	22	2 3
23	Vanidium	V	23	5, 4
24	Chromium	Cr	24	3
25	Manganese	Mn	25	2, 4, 7

26	Iron	Fe	26	2 3
27	Cobalt	Co	27	2 , 3
28	Nickle	Ni	28	1 2
29	Copper	Cu	29	1 , 2
30	Zinc	Zn	30	2

1.3.8 What is Chemical formula?

The chemical formula represents the symbol of elements and ratios of elements to one another in a compound.

Chemical formula tells us number of atoms of each element in a compound with symbols.

For Example: Chemical formula of water is H_2O which indicates that 2 atoms of hydrogen combines with 1 atom of oxygen, or Chemical formula of ammonia NH_3 shows that one nitrogen atom combines with 3 atoms of hydrogen.

1.3.9 Compounds:

The Compound is a substance formed when two or more elements are chemically bonded together in a fixed ratio by mass, As a result a new entirely different properties possessing substance is formed.

The type of bonds holding elements may be ionic bond or covalent bond.

For Example: NaCl, $CuSO_4$, KBr are ionic compounds and H_2O, CH_4, H_2SO_4 are covalent compounds.

Table 1.5 Some Common Compounds with their Formula

Compounds	Chemical Formula
Water	H_2O
Silicon dioxide(sand)	SiO_2
Sodium hydroxide (caustic soda)	NaOH
Sodium chloride(common salt)	NaCl
Sodium carbonate(washing soda)	$Na_2CO_3.10H_2O$
Calcium carbonate(lime stone)	$CaCO_3$
Sugar	$C_{12}H_{22}O_{11}$
Ammonia	NH_3
Sulphuric acid	H_2SO_4
Calcium oxide	CaO

1.3.10 Mixture

When two or more than two elements or compounds physically combined without any fixed ratio is known as Mixture. The component substances retain their chemical properties. Mixtures can be separated again by physical methods, as Filtration, Evaporation, Distillation and Crystallization. There are two main types of mixtures, which are shown in figure 1.5, Homogeneous mixture and Heterogeneous mixture. In a homogenous mixture all the substances are evenly distributed throughout the mixture (Salt water, air, blood). In a heterogeneous mixture the substances are not evenly distributed (chocolate chip cookies, pizza, rocks)

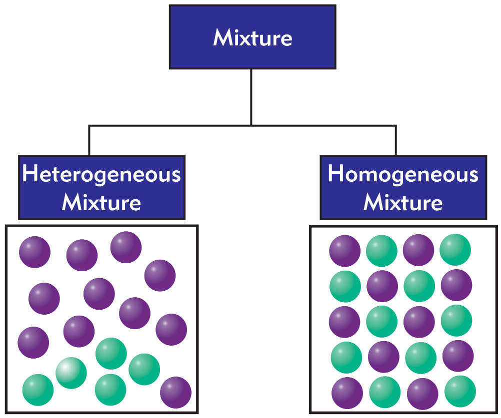


Fig 1.5 Types of Mixture

Table 1.6 Difference between Element, Compound, Mixture

Element	Compound	Mixture
Element is a substance made up of same atoms, and discovered naturally.	Compound is formed by a chemical combination of atoms of the elements.	Mixture formed by the simple mixing of the substances.
Element shows unique properties due to similarity of atoms.	Constituent of compound lose their identity and form a new substance with new properties.	Constituents of mixture retain their properties in mixture.
Atoms of elements have same atomic number.	Compounds have fixed composition by mass.	Mixtures have no fixed composition by mass.
Element cannot be decomposed in to simple substances by ordinary means.	Components cannot be separated by physical means.	The components can be separated by physical means.

Element represented by symbols, which are abbreviations for the names of elements.	Every compound represented by chemical formula.	It consists of two or more components and does not show any chemical formula.
Elements are homogenous	Compounds have homogenous composition.	Mixtures have homogenous as well as heterogeneous composition.
As the atomic number increases in elements melting points increases.	Compounds have sharp and fixed melting points	Mixtures do not have sharp and fixed melting points.

Test Yourself

- How you can differentiate between matter and substance?
- Which elements do the following compounds contains?
  Washing soda, sugar, sand, caustic soda
- Identify mixture, element or compound from the following?
  Table salt, ice cream, blood, silicon, coca cola, tin, zinc, water, sulphur

1.3.11 Relative Atomic Mass and Atomic Mass Unit

Relative Atomic mass:

The Relative atomic mass of an atom is the average mass of naturally occurring isotopes, compared with $^1/_{12}{}^{th}$ mass of one atom of carbon (C-12)

$$A_r = \frac{\text{Average mass of one atom of the element}}{\frac{1}{12} \times \text{the mass of one atom of carbon} - 12}$$

The unit of relative atomic mass is atomic mass unit, with symbol a.m.u .

$$1\text{a.m.u} = 1.66 \times 10^{-24}\text{ gram}$$

1.3.12 Empirical Formula and Molecular Formula

The compounds are represented by Chemical Formula as elements are represented by symbols. Chemical formula are of two types Empirical Formula and Molecular Formula.

Empirical Formula

The formula showing minimum relative numbers of each type of atoms in a molecule is called Empirical Formula.

- Empirical Formula shows simplest ratio of each atoms present in a molecule.
- Empirical Formula does not show the actual number of atoms in the molecule.
- Empirical Formula tells us the type of element present in it.

For Example:

(1) Benzene has molecular formula C_6H_6. The simplest ratio of hydrogen and carbon is 1:1, so the empirical formula becomes CH.

(2) Glucose has molecular formula $C_6H_{12}O_6$. It shows the ratio as follows

$$C_6H_{12}O_6$$
$$6:12:6$$
$$1:2:1$$
$$C:H_2:O$$

So the empirical formula of glucose is CH_2O and has simple ratio 1:2:1 of atoms in molecule of glucose.

Molecular Formula

The Molecular formula is the formula which shows actual number of atoms of each element present in a molecule.

- Molecular formula is derived from empirical formula.
- Molecular formula Mass calculated by adding atomic masses of its atoms.
- Molecular formula of a compound may be same or multiple of empirical formula.

For Example: Molecular Formula of benzene is C_6H_6, which have six carbon and six Hydrogen, molecular formula is an integral multiple (1,2,3 etc.) of the empirical Formula.

Molecular Formula = (Empirical Formula)n where n=1,2,3,etc

Table 1.7 Some Compounds with their Empirical and Molecular Formula

Compound	Empirical Formula	Molecular Formula
Carbon dioxide	CO_2	CO_2
Glucose	CH_2O	$C_6H_{12}O_6$
Hydrogen peroxide	HO	H_2O_2
Benzene	CH	C_6H_6
Acetic Acid	CH_2O	CH_3COOH

1.3.13 Atomic Number and Atomic Mass

The Atomic Number is number of protons present in the nucleus of atom of any Element. It is represented by symbol Z. All atoms of an element have same atomic number due to the presence of same number of protons. For example all oxygen (O_8) atoms have 8 number of protons due to this atomic number is 8 (Z=8).

The Atomic Mass is sum of number of protons and neutrons present in the nucleus of atom of any element. It is represented by symbol A and calculated by A=Z+n where n is number of neutrons. For example nitrogen atom have 7 number of protons and 7 number of electrons then Atomic mass of nitrogen is 14 (A=7+7=14).

Example 1.1: If any element have number of protons 11 and number of neutrons 12, find out its atomic number and atomic mass?

Data:

Number of protons = 11

Number of neutrons = 12

Z = ?

A = ?

As we know atomic number Z is number of protons due to this

Atomic number Z = 11

Atomic mass is A = Z+n

A = 11+12

A = 23

Example 1.2: How many number of protons and neutrons are there in an atom having A= 40 and Z= 20?

Data:

A = 40

Z = 20

Number of protons?

Number of neutrons?

As Number of protons is Z = 20

Number of neutrons = A - Z

= 40 - 20

= 20

1.3.14 Molecular Mass and Formula Mass

Molecular Mass:

The Molecular Mass is the sum of atomic masses of all the atoms present in one molecule of a substance. For example molecular mass of CO_2 is 44 a.m.u and H_2O is 18 a.m.u .

Example 1.3: Calculate the molecular mass of HNO_3

Solution

Atomic mass of H	= 1 a.m.u
Atomic mass of N	= 14 a.m.u
Atomic mass of O	= 16 a.m.u
Molecular mass	= 1(At. Mass of H) +1(At. Mass of N) +3(At. Mass of O)
	= 1 + 14 +3(16)
	= 1 + 14 + 48
	= 63 a.m.u

Formula Mass:

The ionic compounds which form three dimensional solid crystals are represented by their formula units. In such cases formula mass calculated by sum of atomic masses of all atoms in formula unit is called Formula Mass of that substance. For example formula mass of sodium chloride is 58.5 a.m.u

Example 1.4: Calculate the Formula mass of $Al_2(SO_4)_3$

Solution

Atomic mass of Al	= 26.98 a.m.u
Atomic mass of S	= 32 a.m.u
Atomic mass of O	= 16 a.m.u
Formula unit	= $Al_2(SO_4)_3$
Formula mass of $Al_2(SO_4)_3$	= 2(26.98) + 3(32) +12(16)
	= 53.96 + 96 + 192
	= 341.96 a.m.u

Test Yourself

- How you can differentiate empirical formula and molecular formula?
- Why formula mass and molecular mass calculated separately while process of calculation is same?

1.4 CHEMICAL SPECIES:

If one molecule is identical to another we can say they are the same chemical species. Chemical species is a chemical entity, such as particular atom, ion or molecule.

1.4.1 Ions (anions, cations), Molecular ions and Free Radicals

Ions (anions, cations): ion is an atom or group of atoms having a charge on it. The charge may be positive or negative. There are two types of ions, cations and anions. The cations are formed when an atom loses electrons form its outer most shells.

For example: Na^+, K^+ are cations. The following equation shows formation of cations.

$$Na \longrightarrow Na^+ + 1e^-$$

An atom or group of atom that has a negative charge on it is called anion. Anion is formed by the gain or addition of electrons to an atom. For example: Cl^- and O^{2-}, following examples shows formation of an anion by addition of electrons to an atom.

$$Cl + 1e^- \longrightarrow Cl^-$$

Molecular ions: when a molecule loses or gains electrons the species formed called molecular ions. Molecular ions also possess positive or negative charge like any ion. If it has negative charge known as anionic molecular ion, if they have positive charge known as cationic molecular ions. For example CH_4^+, So_4^{2-} etc.

Free Radicals: Free radicals are atoms and group of atoms having number of unpaired electrons. It is represented by putting a dot over the symbol of an element.

For example: H°, Cl° ,H_3C°

Free radicals are formed when homolytic breakage of bond between two atoms takes place by the absorption of heat or light energy. Free radical is very reactive chemical species.

With the above mentioned definitions of ions, molecular ions and free Radicals. Question arises that what is difference between Atom and Ions, Molecule and Molecular ion. Even what is difference between ion and free Radical? Let's discuss it one by one.

Table 1.8 Difference between Atom and Ions

Atom	Ion
Atom is the smallest particle of an element.	Ion is the smallest unit of ionic compound
Atom can or can not exist independently and take part in chemical reaction.	Ion can not exist independently and surrounded by oppositely charged ions.
Atom is electrically neutral.	Ion has negative or positive charge.

Table 1.9 Difference between Molecule and Molecular ion

Molecule	Molecular Ion
Molecule is the smallest particle in a chemical element or compound that has chemical properties of that element or compound.	Molecular ion formed by gain and lose of electrons by a molecule.
Molecule is always neutral.	Molecular ion have positive or negative charge.
Molecule is stable unit.	Molecular ion is reactive species.
Molecule is formed by the combination of atoms.	Molecular ion formed by the ionization of a molecule.

Table 1.10 Difference between ion and Free Radicals

Ion	Free Radicals
Ions are atoms which have positive or negative charge.	Free Radicals are atoms with odd number of unpaired electrons.
Ions exist in crystals and solutions.	Free Radicals exist in air and solutions.
Ion are not affected by the presence of light.	Free Radicals are affected by the presence of light.

1.4.2 Molecule and Types of Molecules:

Molecule:
Molecule is chemical combination of atoms
Molecule is smallest unit of a substance.
Molecule shows properties of substance.
Molecule can exist independently.

(1) Mono atomic Molecule	(2) Di atomic Molecule	(3) Tri atomic Molecule	(4) Poly atomic Molecule	(5) Homo atomic Molecule	(6) Hetero Molecule
Molecule consists of one atom.e.g. helium (He), neon(Ne), argon(Ar)	Molecule consists of two atoms. e.g Hydrogen (H_2) Oxygen (O_2), Chlorine (Cl_2)	Molecule consists of three atoms.e.g. H_2O, CO_2	Molecule consists of many atoms.e.g. CH_4,H_2SO_4 ,$C_6H_{12}O_6$	Molecule consists of same type of atoms.e.g H_2,O_3,P_4,S_8	Molecule consists of different type of atoms.e.g CO_2, H_2O, NH_3

Test Yourself

- Identify the cations, anions, free Radical, molecular ion, molecule from the following.
  O_2 , H^-, N_2 , Cl_2 ,CO_3^{2-}, H_2O , Br^- , H_2 , H_3C° Na^+
- Justify the classification of molecules?

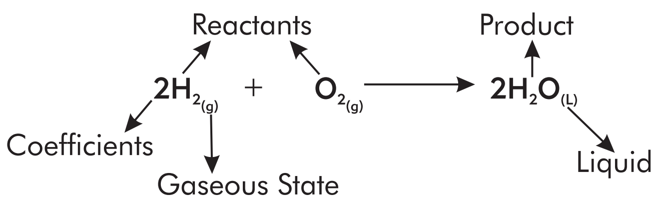
1.5 CHEMICAL EQUATION AND BALANCING CHEMICAL EQUATION

1.5.1 Chemical Equation:

Chemical equation is short hand method of describing the chemical reaction in terms of symbols and formulae of substances.

- The starting substances are known as reactants and always written on the left hand side of arrow,
- The substances which are formed due to reactions of reactants are known as products and written on the right side of arrow.
- The reactants and products are separated from one another by using ($\rightarrow$) single arrow or ($\rightleftharpoons$) double arrow depending on type of reaction.
- The number written in front of formula is called coefficient which shows number of molecules of that reactant or product.
- The expression (s),(g) and(l) shows the state i.e. solid ,gaseous and liquid of reactants and products.
- The expression (aq) expresses that substance is in the form of solution.
  Similarly if a catalyst is used then this catalyst is shown over the arrow.

For example : when two molecules of hydrogen and one molecule of oxygen react then two molecules of water are formed instead of writing the full names of reactants and products, chemists show this reaction as follows in form of equation.



1.5.2 Balancing of Chemical equation

The chemical equation must be balanced in order to fulfill Law of conservation of mass. Mostly chemical equations can be balanced by inspection method (trial and error method). We can balance the equation by following steps.

1.Write the correct formula of all reactants on the left side and products on the Right side of an arrow.

2.Balance the number of atoms on each side.

3.If the number of atoms may appears more or less than other side, balance the equation by inspection method. Multiply the coefficient with formula to make the number of atoms same on the both (reactants and products) sides of equation.

4. The covalent molecules of hyderogen, nitrogen and chlorine exist as diatomic molecules. e.g H_2, O_2, N_2 and Cl_2. We must write them as diatomic molecule rather than isolated atoms in chemical equation.

5. Finally check the equation to be sure that number and kind of atom are same on the reactant and product side. If yes now equation is balanced.

For Example: Let us consider, in laboratory oxygen (O_2) gas is prepared by heating potassium chlorate ($KClO_3$). The products are potassium chloride (KCl) and oxygen (O_2) gas.

Now balance the equation step wise.

Step no1: Write correct formula of all reactants on left side and product on right side of an equation.

$$KClO_{3\,(s)} \longrightarrow KCl_{(s)} + O_{2\,(g)}$$

Step no2: Balance the number of atoms on each side.

Reactants → Products

K (1) → K (1)

Cl (1) → Cl (1)

O (3) → O (2)

We see that K and Cl elements have same number of atoms on both sides of equation but O is not balance because three atoms on reactant side and two atoms on product side.

Step no3: Now multiply the formula ($KClO_3$) with co efficient 2 on reactant side and 3 in front of oxygen on product side to balance the oxygen atoms.

$2KClO_{3\,(s)}$ $KCl_{(s)} + 3O_{2\,(g)}$

Reactants → Products

K (2) → K (1)

Cl (2) → Cl (1)

O (6) → O (6)

Step no4: Now again check and balance the equation by placing 2 in front of KCl on product side.

$$2KClO_{3\,(s)} \longrightarrow 2KCl_{(s)} + 3O_{2\,(g)}$$

Reactants		Products
K (2)	$\longrightarrow$	K (2)
Cl (2)	$\longrightarrow$	Cl (2)
O (6)	$\longrightarrow$	O (6)

Now this chemical equation is balanced.

Test Yourself

- Balance the following equation with coefficient 4 in front of $KClO_3$ on reactant side and 4 in front of KCl on product side.

  $KClO_{3\,(S)} \longrightarrow KCl_{(S)} + O_{2(g)}$
- Balance the following equation.

  $CaCO_3 + HCl \longrightarrow CaCl_2 + H_2O + CO_2$

1.6 MOLE AND AVOGADRO'S NUMBER

1.6.1 Gram Atomic Mass, Gram Molecular Mass, Gram Formula Mass

As we have discussed before that all substances are made up of atoms, molecules or formula units.

The mass of atom is atomic mass, mass of molecule is molecular mass and mass of formula unit is formula mass. All of these masses are expressed in a.m.u. When these masses are expressed in Gram they are termed as Gram atomic mass, Gram molecular mass and Gram formula mass.

Gram Atomic Mass: The atomic mass of an element expressed in gram is called gram atomic mass. It is also called 1 mole.

1 gram atomic mass of oxygen = 16.00g = 1 mole of oxygen atom

1 gram atomic mass of carbon = 12.00g = 1 mole of carbon atom

1 gram atomic mass of nitrogen = 14.00g = 1 mole of nitrogen atom

Its means 1 gram atomic mass of different elements has different masses.

Gram Molecular Mass: The molecular mass of an element or a compound expressed in gram is called gram molecular mass. It is also called 1 mole.

1 gram molecular mass of oxygen (O_2) = 32.00g = 1 mole of oxygen molecule

1 gram molecular mass of water (H_20) = 18.00g = 1 mole of water

1 gram molecular mass of ethanol C_2H_5OH = 46.00g = 1 mole of ethanol

Gram Formula Mass: The formula mass of an ionic compound expressed in grams is called gram formula mass. It is also called 1 mole.

1 gram formula of NaCl = 58.5g = 1 mole of sodium chloride

1gram formula mass of $CaCO_3$ = 100g = 1 mole of calcium carbonate

1.6.2 Mole:

The atomic mass, molecular mass and formula mass of a substance expressed in grams is known as mole. A mole is defined as "amount of substance containing particles equal to the avogadro's number 6.02×10^{23}.

Gram atomic mass and a.m.u

Gram atomic mass of carbon atom = 12g

Atomic mass of one atom of C=12a.m.u

Thus

Gram Atomic mass of carbon is 12gram = 1 mole of carbon atom

Gram Molecular mass of H_2SO_4 is 98gram = 1mole of H_2SO_4

The relationship between mole and mass can be expressed as

$$\text{Number of moles} = \frac{\text{Given mass of a substance}}{\text{Molar mass of the substance}}$$

Or

$$\text{Mass of substance (gm)} = \text{Number of moles} \times \text{Molar mass}$$

Example 1.5: Calculate the number of moles in 40g of Na.

Solution

Given mass of Na =40g

Molecular mass of Na =23 a.m.u

Number of moles =?

$$\text{Number of moles} = \frac{\text{Given mass of a substance}}{\text{Molar mass of the substance}}$$

$$\text{Number of moles of Na} = \frac{40}{23} = 1.73 \text{ moles of Na}$$

Example 1.6: What is the mass of 4 moles of CO_2?

Solution

Number of moles of CO_2 = 4 moles

Formula mass of CO_2 = 44 a.m.u

mass of CO_2 = ?

Mass of CO_2 = number of moles of CO_2 $\times$ formula mass of CO_2

= 4 x 44 = 176 gm

1.6.3 Avogadro's Number

Avogadro an Italian scientist, calculated the number of atoms, molecules or ions present in one mole. The value is found to be 6.02×10^{23}. This value is represented by N_A and is called as Avogadro's number.

For example: 1 mole of O_2 molecule = 32 g

So 32gm of O_2 will contain 6.02×10^{23} molecules

Similarly 1 mole of NaCl =(23+35.5)= 58.5g of NaCl

$=6.02\times10^{23}\ Na^{+} + 6.02\times10^{23}\ Cl^{-}$

Example 1.7: Calculate the number of atoms present in 9.2gm of Calcium (Ca).

Solution:

Atomic mass of Calcium (Ca) = 40

1g atomic weight of Calcium = 40gm

40g of Calcium contains $=6.02\times10^{23}$ atoms of Calcium

By using the formula

$$\text{Number of atoms} = \frac{N_A \times \text{Given Mass in g}}{\text{Atomic mass}}$$

$$\text{Number of atoms} = \frac{6.02\times10^{23}\times 9.2}{40}$$

$$= 1.384\times10^{23}\ \text{atoms of Ca}$$

Example 1.8: Calculate the number of moles, number of molecules present in 8g of $C_6H_{12}O_6$?

Solution

Molecular weight of glucose ($C_6H_{12}O_6$) = (6x12) + (12x1) + (6x16) = 180

Given Mass of $C_6H_{12}O_6$ = 8gm

$$\text{Number of moles} = \frac{8}{180} = 0.04 \text{ mole}$$

$$\begin{aligned}\text{Number of molecules} &= \text{Number of moles} \times N_A \\ &= 0.04 \times 6.02 \times 10^{23} \\ &= 0.240 \times 10^{23} \\ &= 2.40 \times 10^{22} \text{ molecules of glucose}\end{aligned}$$

Test Yourself

- Prove that Avogadro's number is related to mole of any substance.
- Calculate the number of moles in 30gm of H_3PO_4.

1.7 CHEMICAL CALCULATIONS

In all type of chemical calculations we calculate number of moles and number of particles of a substance. These calculations are based on mole. In the sequence of calculation first moles are calculated then number of particles.

1.7.1 Mole-Mass Calculation

In this calculation we calculate number of moles of a substance with the help of following equation.

$$\text{Number of Moles} = \frac{\text{Given mass of substance}}{\text{Molar mass of substance}}$$

We can calculate mass of a substance with the given moles of substance with following equation.

$$\text{Mass of substance} = \text{Number of moles} \times \text{Molar Mass}$$

Example 1.9: A coin of silver (Ag) having 8.5 gm weight. Calculate the number of moles of silver in coin?

Solution

The mass is converted to number of moles by the following equation:

$$\text{Number of Moles} = \frac{\text{Given mass of substance}}{\text{Molar mass of substance}}$$

$$= \frac{8.5}{107}$$

$$= 0.07 \text{ moles of silver in 8.5 gm silver coin}$$

1.7.2 Mole-Particle Calculation

In this calculation we calculate number of moles of a substance in the given number of Particles (atom, molecules or formula unit).

$$\text{Number of Moles} = \frac{\text{Given number of particles}}{\text{Avogadro's number}} = \frac{\text{Given number of particles}}{6.02\text{x}10^{23}}$$

We can calculate number of particles as

$$\text{Number of particles} = \text{Number of moles x } 6.02\text{x}10^{23}$$

Example 1.10: Calculate the number of moles and number of molecules present in 10gm of H_2SO_4?

Solution

The given mass of H_2SO_4 = 10gm

Molar mass of H_2SO_4 = 98.0gm

$$\text{Number of Moles of } H_2SO_4 = \frac{\text{Given mass of substance}}{\text{Molar mass of substance}} = \frac{10}{98} = 0.10 \text{ mole}$$

$$\text{Number Of molecules} = \text{Number of moles x Avogadro's number}$$

$$= 0.10 \text{ x } 6.02\text{x}10^{23}$$

$$= 0.602 \text{ x}10^{23}$$

$$= 6.02 \text{ x}10^{22} \text{ molecules of } H_2SO_4 \text{ in 10g.}$$

1.7.3 Mole Volume Calculation

The mole quantities of gases can be expressed in terms of volume, According to Avogadro, one gram mole of any gas at STP occupies volume of 22.4dm^3 (where standard temperature is O$^{\circ}$C and standard pressures is 1 atm)

Example 1.11: How many liters of carbon dioxide would be produced if 0.450 mole of carbon monoxide reacts with excess oxygen at STP.

Solution:

The equation for the reaction is

$2CO\,(g) + O_2\,(g) \longrightarrow 2CO_2(g)$

$x_1 = 0.450$ moles $\qquad$ x_2 liters = ?

2 moles of CO $\qquad$ 2 moles of Co_2

So,

Step 1 $\quad \dfrac{0.450}{2} = \dfrac{x_2}{2} \longrightarrow x_2 = \dfrac{0.450 \times 2}{2} = \dfrac{0.450}{\text{moles}}$

1 mole of gas at STP means 0°C temperature, 1 atm pressure and occupied volume $22.4dm^3$.

Step 2 $\quad$ Volume in litre/dm^3 = moles × $22.4dm^3$

$= 0.45 \times 22.4 = 10.08$ litre of CO_2

So 10.08 liter of CO_2 would be produced when 0.450 mole of carbon monoxide reacts with excess oxygen at STP.

Summary

- Chemistry is the branch of science which deals with the properties, composition and Structure of matter. Chemistry also deals with the changes involved in the matter.
- Chemistry is everywhere in our environment and serving the humanity day and night. Due to its wide scope Chemistry is divided into physical chemistry, organic chemistry, inorganic chemistry, biochemistry, industrial chemistry, nuclear chemistry, environmental chemistry, analytical chemistry, medicinal chemistry, quantum chemistry, green chemistry.
- Matter is simply defined as anything that has mass and occupies Space.it is found in three common states solid, liquid and gas. The plasma is also considered as fourth state of matter. The different states of matter are due to difference of energy in increasing order.
- Matter is made up of smallest particles which are known as Atom. Atoms are the basic units of matter and define Structure of elements. Now it discovered that atoms are made up of three particles: protons, neutrons and electrons.
- A molecule is the smallest particle in a chemical element or Compound that has the chemical properties of that element or Compound. Molecules are made up of atoms that are held together by chemical bonds. These bonds form as a result of the Sharing or exchange of electrons among atoms. Molecules are Mono, di and poly atomic molecules.

- A piece of matter in pure form is termed as a substance. Every substance has a fixed composition and specific properties. Every substance has physical and chemical properties.
- An Element is a substance made up of same type of atoms having same atomic number and cannot be decomposed into Simple substances by ordinary chemical reaction.
- Elements occur in nature in free or combined form in solid, Liquid and gases states.

  Now 118 elements have been discovered.
- Symbol formula is an abbreviation to represent the name of element. A symbol is taken from the name of that element in English, Latin and Greek. If it is one letter, it will be capital as H for Hydrogen, C for carbon, S for Sulphur and N for Nitrogen etc. in case of two letters symbol, only first letter is capital as Na for sodium, Cr for Chromium, He for Helium and Zn for Zinc.
- When two or more than two elements or compounds physically combined without any fixed ratio is known as Mixture. The component substances retain their chemical properties. Mixtures can be separated again by physical methods, as Filtration, Evaporation, Distillation and crystallization.
- The Atomic Number is number of protons present in the nucleus of atom of any Element. It is represented by symbol Z. All atoms of an element have same atomic number due to the presence of same number of protons.
- The Atomic Mass is sum of number of protons and neutrons present in the nucleus of atom of any element.it represented by symbol A and calculated by $A=Z+n$ where n is number of neutrons.
- The atomic mass of an element expressed in gram is called gram atomic mass. It is also called 1 mole.
- The molecular mass of an element or a compound expressed in gram is called gram molecular mass. It is also called 1 mole.
- The formula mass of an ionic compound expressed in grams is called gram formula mass. It is also called 1 mole.
- The atomic mass, molecular mass and formula mass of a substance expressed in grams is known as mole.
- Avogadro an Italian scientist, calculated the number of atoms, molecules or ions present in one mole. The value is found to be 6.02×10^{23} . This value is represented by N_A and is called as Avogadro's number.

Exercise

SECTION- A: MULTIPLE CHOICE QUESTIONS

Tick Mark (✓) the correct answer

1. The branch of Chemistry which deals with hydrocarbons:

(a) Industrial chemistry (b) Inorganic chemistry

(c) Organic chemistry (d) Physical chemistry

2. The atomic mass of an element expressed in gram is:

(a) Gram molecular mass (b) Gram atomic mass

(c) Gram formula mass (d) Mole

3. Which of the following can be separated by physical means?

(a) Mixture (b) Element

(c) Compound (d) Substance

4 The molar mass of $H_2SO_4$is:

(a) 98 a.m.u (b) 9.8gm

(c) 98gm (d) 9.8 a.m.u

5. The Molecule consist of two atoms is :

(a) Monoatomic molecule (b) Polyatomic molecule

(c) Hetero atomic molecule (d) Di atomic molecule

6. A formula that indicates actual number and type of atoms in a molecule is called:

(a) Chemical formula (b) Empirical formula

(c) Molecular formula (d) Formula mass

7. Ethyl alcohol was prepared by:

(a) Ibne-Sina (b) Al-Razi

(c) Al-Beruni (d) Jaber bin-Hayan

8. which of the following is a homo atomic molecule:

(a) H_2 (b) NH_3

(c) H_2O (d) CO_2

9. The Empirical formula of hydrogen peroxide (H_2O_2) is:
   (a) H_2O_2 (b) HO
   (c) OH_2 (d) O_2H_2
10. A piece of matter in pure form is termed as:
   (a) Redical (b) Mixture
   (c) Compound (d) substance

SECTION- B: SHORT QUESTIONS:

1. Differentiate between the physical and analytical chemistry?
2. Write down the classification of molecule?
3. Identify the differences among the following:
   (a) Atom and Ion
   (b) Molecule and Molecular ion
   (c) Compound and Mixture
4. Define the following terms:
   (a) Gram atomic mass
   (b) Gram molecular mass
   (c) Gram formula mass
5. Write down the empirical and molecular formula of the following?
   Sulphuric acid, Carbon dioxide, Glucose, Benzene
6. What is Free Radical?
7. Describe relationship between empirical and molecular formula? Explain with examples.
8. Explain why hydrogen and oxygen are considered as element whereas water is compound?

SECTION- C: DETAILED QUESTIONS:

1. What do you mean chemical species, explain ion, molecular ion and free radical?
2. Write down the applications of chemistry in daily life?
3. Explain in detail empirical and molecular formula?
4. Explain the steps for balancing the equation?
5. Name the branches of chemistry and discuss any five branches?

SECTION- D: Numerical

.(1) Balance the following equations by inspection method:

(a) $NH_3 + O_2 \longrightarrow NO + H_2O$

(b) $KNO_3 \longrightarrow KNO_2 + O_2$

(c) $Ca + H_2O \longrightarrow Ca(OH)_2 + H_2$

(d) $NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$

(e) $CO + O_2 \longrightarrow CO_2$

(2) Calculate the formula mass (a.m.u) of the following?

Al_2O_3 , $MgCl_2$, $NaCl$, KNO_3

(3) Calculate the molecular mass (a.m.u) of the following?

C_2H_5OH, H_2O, NH_3, CO_2

(4) How many moles are present in 40 gm of H_2SO_4?

(5) Calculate the number of moles and number of molecules present in the following?

(a) 16 g of H_2CO_3 (b) 20g of $C_6H_{12}O_6$

Chapter 2

ATOMIC STRUCTURE

Time Allocation

Teaching periods = 14
Assessment period = 4
Weightage = 14

Major Concepts

2.1 Discovery of Sub Atomic Particles, electron, proton and neutron.

2.2 Theories and Experiments Related to Atomic Structure.

2.3 Modern Theories of Atomic Structure.

2.4 Electronic Configuration.

2.5 Isotopes and their common application.

STUDENTS LEARNING OUT COMES (SLO'S)

Students will be able to:

- Describe the discovery of electron, proton and neutron.
- Define Atomic Number (Z) and Mass Number (A) in term of number of proton and/or neutron.
- Describe the contributions Rutherford made to the development of the atomic theory.
- Explain how Bohr's atomic model is different.
- Define Modern theories of Atomic Structure(De Broglie Hypothesis & Schrodinger atomic model)
- Describe the presence of sub shells in a shell.
- Distinguish between Shells and Sub shells.
- Write Electronic Configuration of the first 18 Elements in the Periodic Table.
- Define and compare isotopes of an Atom.
- Discuss the properties of the isotopes of the H, C, Cl and U.
- Draw the structure of different isotopes from mass number and atomic number.
- State the importance and uses of the isotopes in various fields of life.

Introduction

The word atom is derived from a Greek word ATOMOS means indivisible, which was first describe by Greek philosopher Democritus. Democritus beleived that all matter consist of very small indivisible particles which are known as atoms. Johan Dalton an English school teacher and chemist suggested the fundamental atomic theory, which explains that all elements are made up of tiny indivisible particles called atoms. Dalton assumed that no particles smaller than atom exist, but by the passage of time new experiments showed that atom is composed of even smaller particles which are known as sub-atomic particles. After that these sub-atomic particles were discovered and named as electron, proton and neutron. We will discuss all these discoveries in this chapter.

Fig 2.1 Democritus

2.1 Discovery of Sub Atomic Particles (Electron, Proton, Neutron) of an Atom

Dalton's atomic theory explains the chemical nature of matter and existence of indivisible atoms, but at the end of 19^{th} century sub-atomic particles were discovered by different scientists. First sub-atomic particle Electron was discovered by M. Farady, William Crooks and J.J. Thomson, second sub-atomic particle Proton was identified by Goldstein and Ernest Rutherford, while third sub-atomic particle Neutron was revealed by Chadwick. All of these findings were milestone in the knowledge of atomic structure which we have now.

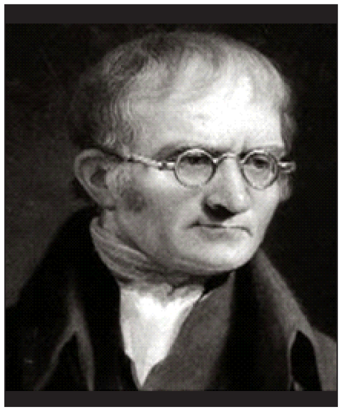


Fig 2.2 john Dalton

2.1.1 Discovery of Electrons

Electron is the lightest particle carrying negative charge in an atom discovered by J.J.Thomson and William crooks.

Fig 2.3 Chadwick


Fig 2.4 J.J.Thomson

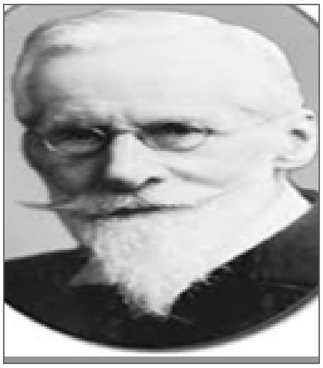


Fig 2.5 William Crooks


Fig 2.6 M.Faraday

The apparatus used for this type of experiment is called discharge tube which consists of glass tube fitted with two metal electrodes connected to a high voltage source and a vacuum pump. Discharge tube inside is evacuated, and electrodes are connected with high voltage source at very low pressure(1mm of Hg), when the high voltage current start passing between electrodes then a streak of bluish light originate and travel in straight line from cathode (-ve electrode) to anode(+ve electrode), Which cause glow at the wall of opposite end. These rays are called cathode rays.

Discovery of Electrons

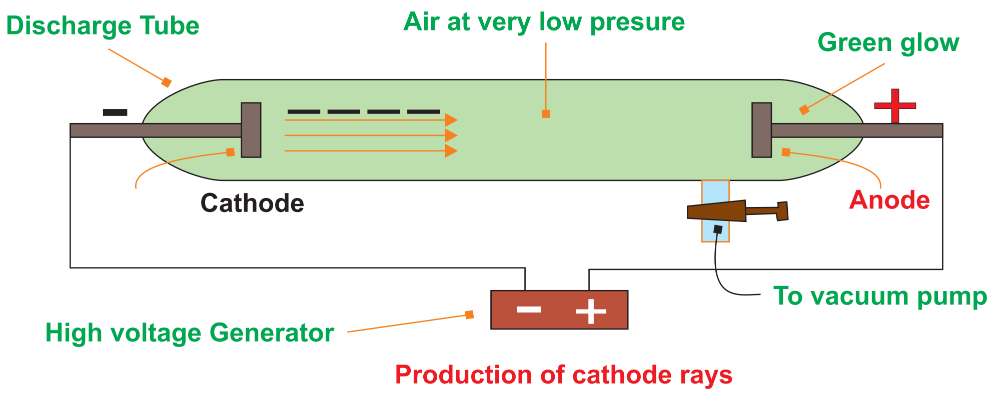


Fig 2.7 William Crooks Discharge Tube

J.J.Thomson justified that these rays were deflected towards positive plate in electric and magnetic field which shows that these rays possess negative charge due to this negative charge, particle was named Electron .These electrons were obtained from the cathode and when cathode material was changed the same phenomenon was observed which proves that electrons are constituent of all matter.

Properties of Cathode Rays (Electrons)

1. They travel in straight line from cathode towards Anode.
2. They produce sharp shadow of an opaque object placed in their path.
3. They have negative charge and bend towards positive plate in electric and magnetic field.
4. These rays when strike with glass or other material and cause the material to glow.
5. The charge to mass ratio (e/m) of cathode particles is 1.7588×10^8 coulomb per gram. This is same for all electrons, regardless of any gas in discharge tube.
6. They can produce mechanical pressure indicating they possess kinetic energy (K.E).

2.1.2 Discovery of Protons

The Proton is positively charged particle discovered by Goldstein in 1886. J.J.Thomson investigated properties of proton in 1897.

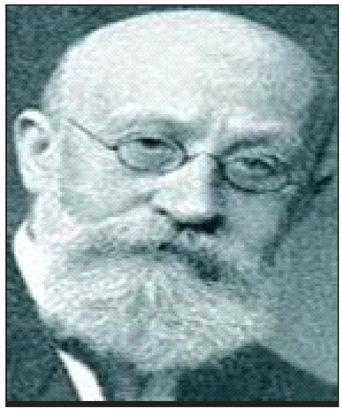


Fig 2.8 Goldstein

Protons were observed in same apparatus of cathode rays tube but with perforated cathode. Goldstein discovered that not only negatively charged cathode rays but positively charged rays are moving in opposite direction through perforated cathode. These positive rays passes through the holes of cathode, where they strike with walls of tube and cause the glow of tube. These rays were named as Canal rays (protons).

Discovery of Protons

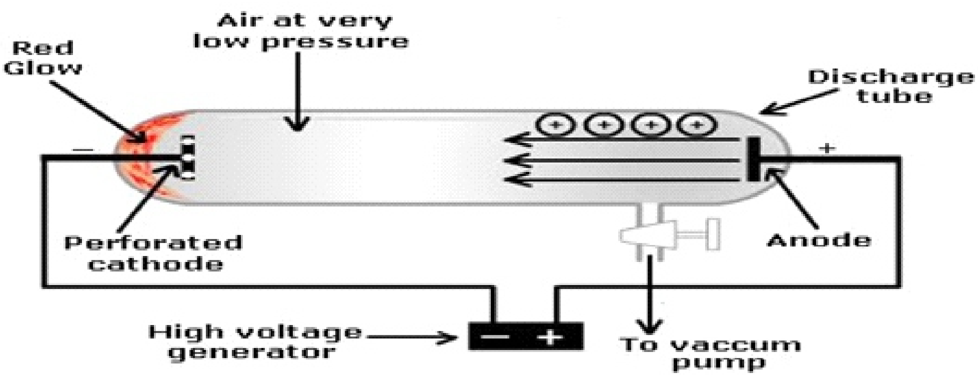


Fig 2.9 Gold Stein Discharge Tube

Remember that canal rays are not emitted by anode, but they are result of striking of electron with residual gas molecules in discharge tube. Electrons ionize the gas molecules as follows.

$$M + e^{-} \longrightarrow M^{+1} + 2\,e^{-}$$

Goldstein justified that atoms are electrically neutral, while electrons carry negative charge. It mean for each electron there must be one equivalent positive charge to neutralize that electron. This particle is called proton and it is a fundamental particle of all Atoms.

Properties of Canal Rays (Protons)

1. They travel in straight line towards Cathode.
2. They produce sharp shadow of object placed in their path.
3. They have positive charge and bend towards negative plate in electric and magnetic field.
4. The charge to mass ratio (e/m) of positive particles is much smaller than electron. It varies according to nature of gas present in tube.
5. The mass of proton is 1836 times more than electron.

2.1.3 Discovery of Neutrons

In 1920 Rutherford predicted that atom must possess another neutral particle with equivalent mass of proton. Different scientists started working on this neutral particle. later on, in 1932 Chadwick successfully discovered Neutron. Chadwick found that when alpha (α)particles bombarded on Beryllium plate some penetrating radiations were given out. Chadwick suggested that these radiations were due to material particle with mass comparable to hydrogen atom but have no charge. These radiations (particle) are called Neutron. It can be expressed in equation as follows.

$${}^{9}_{4}Be + {}^{4}_{2}He \longrightarrow {}^{12}_{6}C + {}_{0}n^{1}$$

The neutron is fundamental part of an Atom, present inside nucleus with proton and is included in atomic mass.

Properties of Neutrons

1. The Neutrons are neutral particles.
2. They have no charge.
3. The mass of neutron is almost equal to that of proton.
4. These particles are most penetrating in matter.

2.1.4 How Atomic Number (Z) and Mass Number (A) are related with number of proton and neutron

As we discussed in discovery of fundamental particles of an atom, that Atom consist of three particles Electron, Proton and Neutron. But if all atoms have same fundamental particles then why the atoms of one element are different from the atoms of another element?

For example: How does an atom of Carbon (C) is different from an atom of Nitrogen (N)?Because all atoms can be identified by their number of protons they contain. Therefore no two elements have the same number of protons.

Atomic Number (Z)

The number of protons inside the nucleus of an atom is called Atomic Number. Atomic number is represented by Z. The elements are identified by their atomic number. Different elements have different atomic numbers because of different number of protons. In neutral atoms number of protons are equal to number of electrons, so the atomic number also indicate total number of electrons outside the nucleus. For example atomic number of Carbon(C) is 6. It mean that each carbon atom has 6 protons and 6 electrons in it.

Atomic number= Z = Number of proton in nucleus = Total number of electron around nucleus

Atomic number (Z) is written as subscript on the left hand side of the chemical symbol e.g ${}_{6}C$. Some other examples are as follows.

$${}_{3}Li, \quad {}_{8}O, \quad {}_{1}H$$

Mass Number (A)

The total sum of protons and neutrons in the nucleus of an atom is called Mass Number. It is also known as "Nucleon number"Mass number represented by A. For example, the sodium (Na) atom has atomic number 11 and mass number 23. It indicates that sodium atom has 11 protons and 12 neutrons. The mass number (A) is written as superscript on left hand side of chemical symbol. e.g ${}^{23}_{11}Na$

$$\text{Mass number} = A = \text{Number of protons (Z)} + \text{Number of neutrons (N)} \quad \text{OR}$$

$$\text{Mass number} \quad A = Z + N$$

$$\text{And} \quad \text{number of neutron} \quad N = A - Z$$

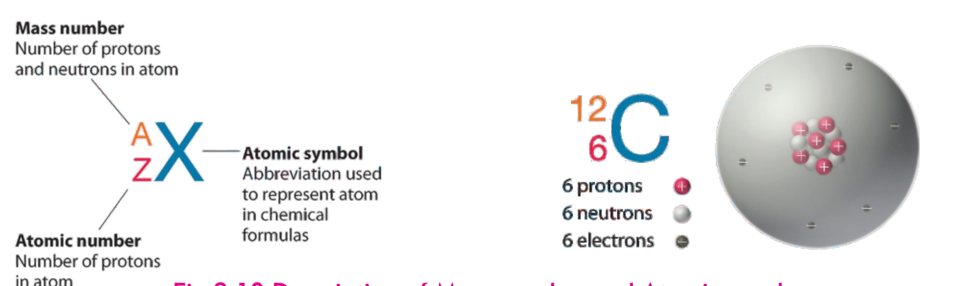


Fig 2.10 Description of Mass number and Atomic number

Test Yourself

a. What is atomic number of an oxygen atom which have 8 Neutrons and 8 protons?

b. Find out mass number of chlorine which have 17 protons and 18 neutrons?

c. How many electrons, protons and neutrons are present in Co ?

d. Do you know any element which have no neutron in its atom?

2.2 Theories and Experiments Related to Atomic Structure

2.2.1 Rutherford Atomic Model

Lord Rutherford in 1911, carried out series of experiments and proposed a new model for the atom.

EXPERIMENT

Rutherford took a thin sheet of gold foil and bombarded it with alpha (α) particles obtained from a radioactive element (Like Polonium). These rays were scattered after passing through the foil and examined on a zinc sulphide (ZnS) screen.

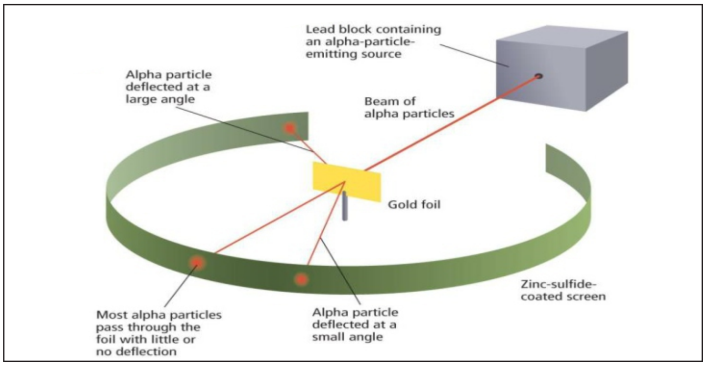


Fig 2.11 Gold Foil experiment

Do you know?

Radioactive element are unstable isotopes that release subatomic particles or energy as they decay.

For example:

Uranium, Radium and Polonium

Observations

1. Most of the particles passed straight and un deflected through the sheet and produced illumination on the zinc sulphide screen.
2. Very few alpha (α) particles undergo small and strong deflection after passing through gold sheet.
3. A very few alpha (α) particles (one out of 8000) retraced their path.

Do you know?

Illumination:

It is the action of supplying or brightening with light The luminous flux per unit area on an intercepting surface at any given point called illumination

Conclusion

1. According to Rutherford an atom consist of two parts nucleus and extra nuclear part.
2. Majority of the alpha particles passed straight line and un-deflected, shows that most volume occupied by atom is empty.
3. Alpha particles are positively charged and their deflection indicates that the centre of atom has a positive charge, which is named as nucleus.
4. The mass is concentrated in the nucleus and the electrons are distributed outside the positively charged nucleus.
5. The electrons are revolving around the nucleus in extra nuclear part called orbits.

Conclusion of Ruther Ford "Gold Foil" experiment

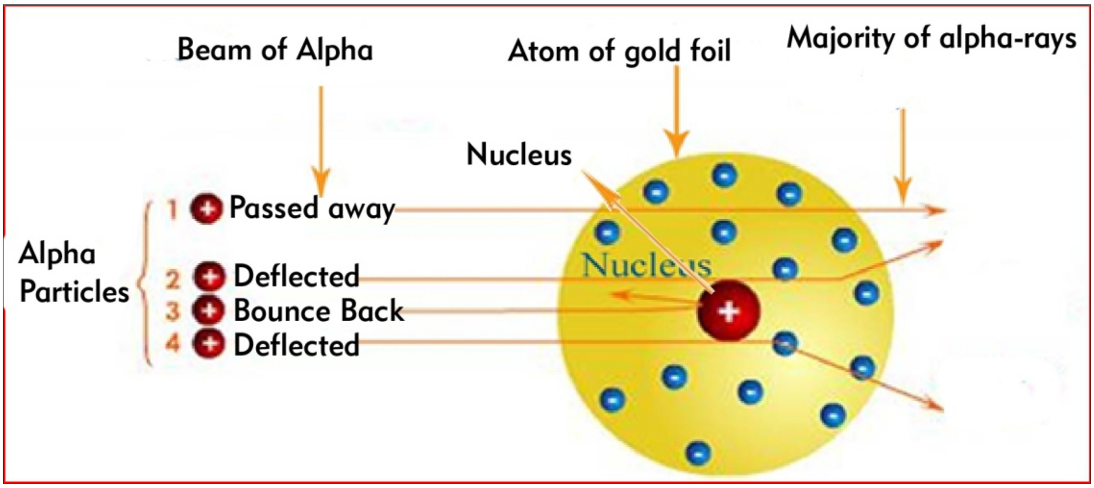


Fig 2.12 Pictorial description of bombardment of alpha particles on gold foil

Rutherford postulates

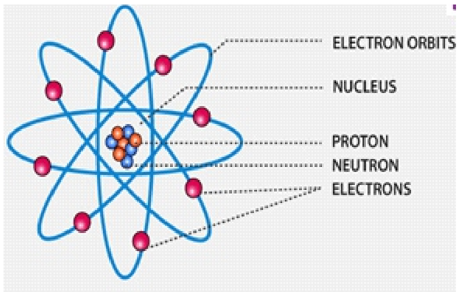


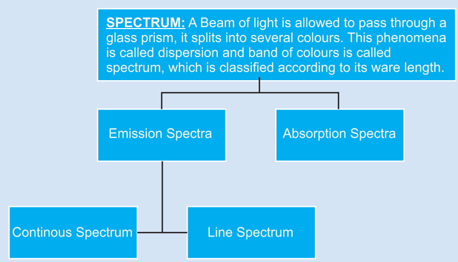
Fig 2.13 Rutherford Atomic Model

- An atom consists of positively charged, dense and very small nucleus containing protons and neutrons. The entire mass is concentrated in the nucleus of an atom.
- The nucleus is surrounded by large empty space which is called extra nuclear part where probability of finding electron is maximum.
- The electrons are revolving around the nucleus in circular paths with high speed (Velocity).
- These circular paths are known as orbits (Shells).
- An atom is electrically neutral because it has equal number of protons and electrons.
- The size of the nucleus is very small as compared to the size of its original atom.

Defects of Rutherford atomic model

1. Rutherford did not explain the stability of an atom.
2. In Rutherford atomic model the negatively charged electrons revolve around the nucleus in circular path and emits energy continuously. Due to continuous loss of energy ultimately it must fall into the nucleus.
3. If the revolving electron emits energy, continuously then there would be a continuous spectrum but in contrast to it we get line spectrum from the atoms of elements.

Do you know?



2.2.2 Neil Bohr's Atomic Model

In 1913 Neil Bohr proposed another atomic model. This atomic model was different in this manner that it showed two folds, first removed the Rutherford atomic model and second explained the line spectrum of Hydrogen atom based on quantum theory of Max Planck.

Do you know?

What is quantum?

A discrete quantity of energy proportion which can exist independently.

Postulates of Neil Bohr's Atomic Model

Neil Bohr proposed the following postulates for atomic structure.

1. The atom has fixed orbits in which negatively charged electrons are revolving around the positively charged nucleus.
2. These orbits possess certain amount of energy which are called shells and named as K, L, M, N shells.
3. The energy levels are represented by an integer (n= 1, 2, 3....) known as quantum number, this quantum range starts from nucleus side, where n=1 is lowest energy level.
4. Electrons are revolving in particular orbits (ground state) continuously, but they do not emit energy.
5. When electron absorbs energy, it jumps from lower energy level (E_1) to higher energy level (E_2) (excited state).
6. When electrons jumps back from higher energy level (E_2) to lower energy level (E_1), it emits energy.
7. The emission or absorption is discontinuous in the form of energy packet called Quantum or Photon.
8. The ΔE difference in energy of higher (E_2) and lower (E_1) energy level.

   $\Delta E = E_2 - E_1$

   $\Delta E = h\nu = 1$ photon

Here h is planks constant, its value is 6.63×10^{-34} Js and ν is a frequency of light.

9. Stationary state were present in those orbits in which angular moment of electron would be integral multiple of $h/2\pi$

   $mvr = nh/2\pi$ (where n = no: of orbits) h=(planks constant) m= (mass of electron)

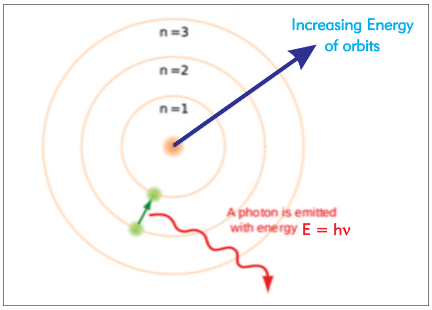


Fig 2.14 Neil Bohr's Atomic model

Limitations of Bohr's Atomic Model :

- Bohr's model of an atom failed to explain the Zeeman Effect (effect of magnetic field on the spectra of atoms).
- It also failed to explain the Stark effect (effect of electric field on the spectra of atoms).
- It deviates the Heisenberg Uncertainty Principle.
- It could not explain the spectra obtained from larger atoms.
- It only explain the monoelectronic species like He^{+}, Li^{+2}, Be^{3+}.

Test Yourself

- Which particles shows mass of an atom?
- Prove Rutherford atomic model based on classical theory and Bohr atomic model based on quantum theory?
- How you can relate living things with chemistry?

2.3 Modern Theories of Atomic Structure

In the year of 1900 Max Planck proposed quantum nature of radiations and energy in a photon wave E=hν as quantum theory. This quantum theory accepted by Albert Einstein in 1905 and proposed relationship between mass and energy to explain photoelectric effect by wave particle duality as $E = mc^2$. In 1913 Neil Bohr continue to use quantization of radiation with angular momentum of electrons. Bohr predicted and explained the line spectrum of Hydrogen atom

2.3.1 de Broglie Hypothesis

In 1923 Lois de Broglie extended the wave particle duality to electron, and proposed a hypothesis that all matter has particle as well as wave nature at the submicroscopic level.

De Broglie combined the Einstein and Planck equations and argued that if

$E=h\nu$ where E = energy, h = plank 'constant, ν= frequency of light

And $E = mc^2$ where E = energy, m = mass, c = speed of light

Then

$$h\nu = mc^2$$

$$\frac{hc}{\lambda} = mc^2 \qquad \left(\nu = \frac{c}{\lambda}\right)$$

or $$\lambda = \frac{h}{mc} \qquad (P = mc)$$

or $$\boxed{\lambda = \frac{h}{P}} \qquad \text{or } P = m\nu$$

The wave nature of a particle is quantified by De Broglie wavelength defined as $\lambda = h/p$ where p is the momentum of the particle.

According to De-Broglie a light, or any other electromagnetic wave, can also exhibit the properties of a particle, similarly a particle should also exhibit the properties of a wave, and those two nature are interchangeable.

DeBroglie wave particles duality hypothesis

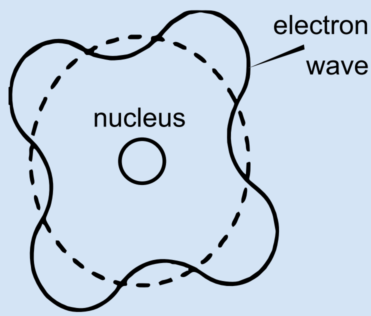


Fig 2.15 De Broglie wave duality hypothesis

2.3.2 Schrodinger Atomic Model

In 1926 Erwin Schrödinger, an Austrian physicist, took the Bohr's atomic model one step forward. Schrödinger used mathematical equations to describe the likelihood of finding an electron in a certain position. This atomic model is known as the quantum mechanical model of the atom.

Schrodinger model is just an improvement of Bohr's atomic model. He took an atom of hydrogen because it has one proton and one electron. He proved mathematically that electron can be find in different position around the nucleus and determined by probability.

- The quantum mechanical model determines that electron can be find in various location around the nucleus. He found electrons are in orbit as an electron cloud.
- Each subshell in an orbit have different shapes which determine the presence of electron.
- Different subshells of orbitals are named as s, p, d and f with different shapes as 's' is spherical and 'p' is dumbbell shaped.
- The numbers and kind of atomic orbitals depends on the energy of shell.

According to quantum mechanical model probability of finding an electron within certain volume of space surrounding the nucleus can be represented as a fuzzy cloud. If the cloud is denser the probability of finding electron is high which are called atomic orbitals. Detail and mathematical derivation will be discussed in next classes.

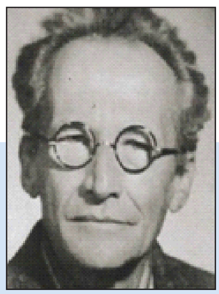


Fig 2.16 Schrodinger

2.4 Electronic Configuration

Before discussing electronic configuration we must understand the concept of shells and subshell.

As we know that nucleus is present in the centre of an atom and around the nucleus electrons are revolving. Now we have to understand how these electrons are revolving around the nucleus. These electrons are revolving around nucleus in different levels according to their potential energy.

2.4.1 Concept of Shell (K, L, M, N,O, P & Q)

The Energy levels or Shell or Orbit are all possible paths on which electrons are revolving around nucleus. Which is shown by 'n'. These shells are named as K, L, M, N, O, P & Q with quantum numbers n= 1, 2, 3, 4, 5, 6 and 7 respectively. These shells have definite amount of energy by means of increasing order as they become away from nucleus.

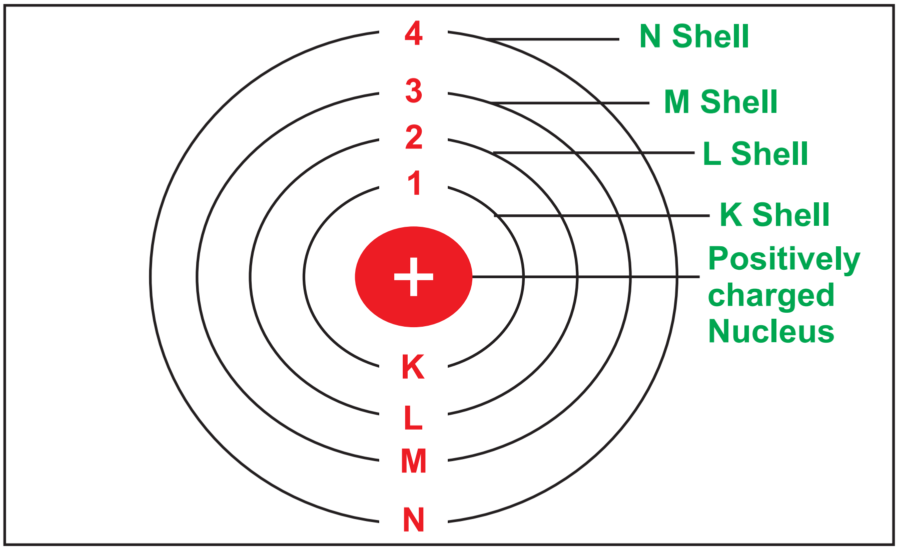


Fig 2.17 Shell (Energy level)

First energy level is K shell and has less energy.

Second energy level is L shell and has more energy than K shell.

Third energy level is M shell and has more energy than K and L shells.

Fourth energy level is N shell and has more energy than K, L and M shells.

Fifth energy level is O shell and has more energy than K, L, M and N shells.

2.4.2 Concept of Sub Shell (s,p,d & f)

When atomic spectra of substances were observed in a high powered spectroscope, it was found that they consist of two or more lines closely packed with each other as discussed in zemen and stark effects. These lines means that electrons in the same shell may differ in energy by small amount. Thus main energy level are divided in to sub energy levels and known as sub shells. When electrons are many in numbers in a shell they show repulsion and main shell splits into subshell which are named as s, p, d and f subshells.

The number of subshells in a shell is according to value of that shell, which are given in table 2.1

Table 2.1 Values of shell and sub shell

Value of 'n'	Shell	Sub shell
1	K	Only s
2	L	s, p
3	M	s. p,d
4	N	s,p, d, f

Do you know?

The atomic spectrum of substance consists of spectral lines. These lines differ in energy by small amount. Energy levels are divided into subshells/ subenergy levels due to repulsion. The shell or orbit split into subshell which are named as s, p, d, and f.

2.4.3 Electronic Configuration of First 18 Elements

Now we can understand that 'the distribution of electrons among the different orbits/shells and subshells according to some rules is known as the electronic configuration of an atom'. Generally, the most stable electronic configuration is represented when an atom is at the ground state with less energy level. Electrons filled in increasing order from lower to higher energy levels as

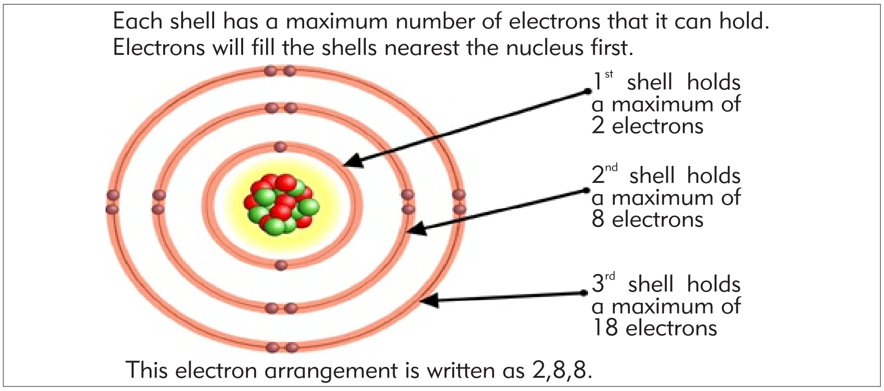


Fig 2.18 Filling of energy level

The maximum number of electrons that can be accommodated in a shell is represented by the formula $2n^2$, where 'n' is the shell number. The distribution of electrons in different orbits are as follows:

K-shell/ 1^{st} orbit (n=1) = $2(1)^2$ = 2
L-shell/ 2^{nd} orbit (n=2) = $2(2)^2$ = 8
M-shell/ 3^{rd} orbit (n=3) = $2(3)^2$ =18
N-shell/4^{th} orbit (n=4) =$2(4)^2$ =32 and so on

There is slight difference in Energy of subshells, the subshell "s" filled first then subshell 'p' and onward. The distribution of maximum electrons in subshells is as follows.

2 electrons in's' subshell

6 electrons in 'p' subshell

10 electrons in 'd' subshell

14 electrons in 'f' subshell

Whenever we write electronic configuration always remember following points.

1. Number of Electrons in an Atom.
2. Arrangement of shells and subshells according to energy levels.
3. Maximum number of electrons for shells and subshells.

Example 2.1: write down electronic configuration of an element which has 8 electrons.

For this element first of all electrons will be filled in K shell which have maximum capacity of 2 electrons, than remaining electrons will be filled in L shell which has maximum capacity of 8 electrons. Now arrangement of electrons will be as:

K L M

2, 6, 0

The above element is Oxygen which has 8 electrons. In writing electronic configuration first two electrons will go into '1s' subshell of K shell which hold two electrons. The next two electrons for oxygen go in the 2s subshell of L shell and remaining four electrons will go into 2p subshell of L shell. Now electronic configuration of oxygen is $1s^2\ 2s^2\ 2p^4$

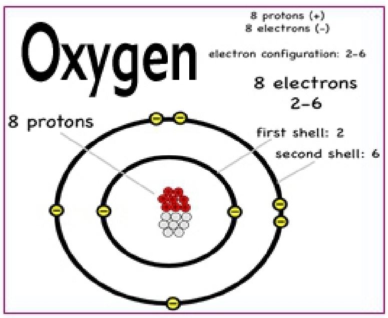


Fig 2.19
Electronic Configuration of Oxygen

The electronic configuration of different subshell of atom is written as $1s^2$,$2s^2$, $2p^6$, $3s^2$, as shown in figure 2.20.

Where coefficient shows number of shell, s,p are subshells and superscript is number of electrons in subshells. The electronic configuration of first 18 elements is given in table 2.2

Order of Filling of Electrons in Subshells

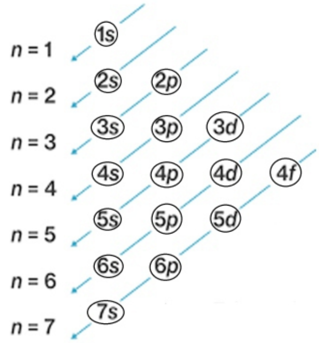


Fig 2.20 Order of filling of electrons in sub shells

Table 2.2 Electronic arrangement of the first 18 elements of the periodic table

Elements	Symbol	Atomic Number (number of electrons)	Electronic Configuration
Hydrogen	H	1	$1s^1$
Helium	He	2	$1s^2$
Lithium	Li	3	$1s^2$, $2s^1$
Beryllium	Be	4	$1s^2$, $2s^2$
Boron	B	5	$1s^2$, $2s^2$, $2p^1$
Carbon	C	6	$1s^2$, $2s^2$, $2p^2$
Nitrogen	N	7	$1s^2$, $2s^2$, $2p^3$
Oxygen	O	8	$1s^2$, $2s^2$, $2p^4$
Fluorine	F	9	$1s^2$, $2s^2$, $2p^5$
Neon	Ne	10	$1s^2$, $2s^2$, $2p^6$
Sodium	Na	11	$1s^2$, $2s^2$, $2p^6$,$3s^1$
Magnesium	Mg	12	$1s^2$, $2s^2$, $2p^6$,$3s^2$
Aluminum	Al	13	$1s^2$, $2s^2$, $2p^6$,$3s^2$,$3p^1$
Silicon	Si	14	$1s^2$, $2s^2$, $2p^6$ $3s^2$,$3p^2$
Phosphorus	P	15	$1s^2$, $2s^2$, $2p^6$,$3s^2$,$3p^3$
Sulphur	S	16	$1s^2$, $2s^2$, $2p^6$,$3s^2$,$3p^4$
Chlorine	Cl	17	$1s^2$, $2s^2$, $2p^6$,$3s^2$,$3p^5$
Argon	Ar	18	$1s^2$, $2s^2$, $2p^6$,$3s^2$,$3p^6$

Test Yourself

- What is maximum number of electrons that can be accommodate in's' subshell?
- How many electrons will be in L shell of an atom having atomic number 11?
- In the distribution of electrons of an atom, which shell filled first and why?
- If both K and L shells of an atom are completely filled, what is the total number of electrons are present in them?

2.5 ISOTOPES AND THEIR COMMON APPLICATION

As we know that atom is composed of three particles electrons, protons and neutrons. In all the atoms of an element number of electrons, protons and number of neutron are same, due to this their atomic number and mass number are same but in few elements some atoms have different atomic number and mass number.

2.5.1 What are Isotopes?

Atoms of the same elements having same atomic number but different Mass number are called isotopes. They have same number of electrons and protons but different number of neutrons. These elements have same chemical properties due to same electronic configuration but different physical properties due to difference in mass number.

2.5.2 Examples of Isotopes

(1) Isotopes of Hydrogen

There are three isotopes of Hydrogen. These are known as Protium, deuterium and tritium as shown in fig 2.21

Protium

Deuterium

Tritium

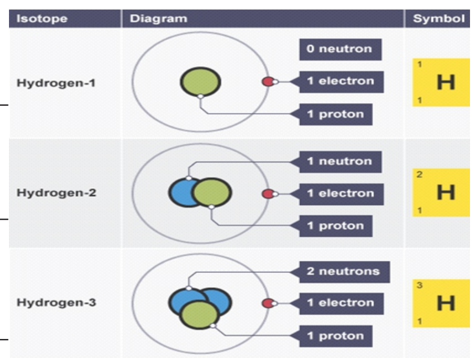


Fig 2.21 Isotopes of Hydrogen

(2) Isotopes of Uranium

There are three common isotopes of uranium with atomic number 92 and mass number 234,235 and 238 respectively, as shown in fig 2.22. The uranium $^{238}_{92}U$ is found 99% in nature.

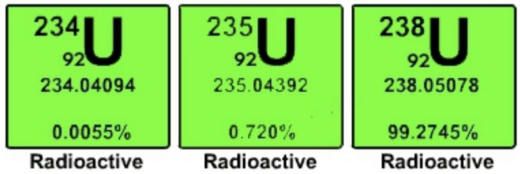


Fig 2.22 Isotopes of Uranium

(3) Isotopes of Carbon

There are two stable isotopes and one radioactive isotope of carbon. Which are shown in fig 2.23.

The carbon 12 contain 6 proton and 6 neutron, Carbon 13 possess 6 proton and 7 neutron, carbon 14 contain 6 proton and 8 neutron. Carbon 12 is the most abundant (98.89%) isotope.

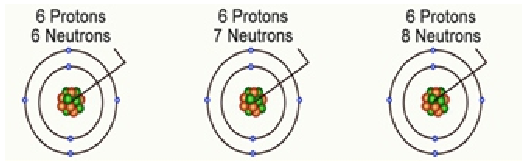


Fig 2.23 Isotopes of Carbon

(4) Isotopes of Chlorine

There are two isotopes of Chlorine with atomic number 17 and mass number 35 and 37. as shown in figure 2.24. Chlorine 35 is 75% and chlorine 37 is 25% abundant in nature.

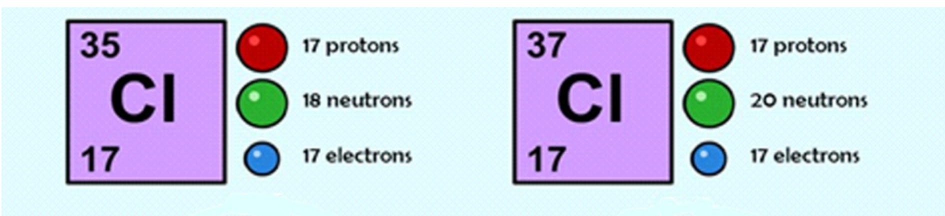


Fig 2.24 Isotopes of Chlorine

Table 2.3 Applications of Isotopes

S. No	Name of Radioactive Isotopes	Fields	Uses
(1)	Phosphorous -32 or strontium -90	Radiotherapy	• Treatment of skin cancer
(2)	Cobalt-60	Radiotherapy	• Treatment of body cancer due to more penetrating power.
(3)	Iodine isotopes	Radiotherapy	• Detestations of thyroid glands in he neck.
(4)	Technetium	Radiotherapy	• To monitor the bone growth in fracture healing.
(5)	Gamma ray of cobalt – 60	Medical instrumentation	• To sterilization of medical instruments and dressings from harmful bacteria.
(6)	Americium -241	Safety measures & industries	• Used in back scatter gauges, smoke detectors fill height detectors and measuring ash content of coal.
(7)	Gold -198 and Technetium - 99	Sewage & liquid waste movement for water pollution	• Tracing factory waste causing ocean pollution • Tracing sand movement in rivers and oceans.
(8)	Uranium -235	Power Generation	• Conversion of water energy from steam to generate electricity.
(9)	Plutonium -238	Medicine	• Used to stimulate a regular heart beat in heart pace maker.
(10)	Carbon -14	Archaeology and Geology	• Used to estimate the age of fossils.

Test Yourself

- Which of the isotopes of hydrogen contains greater number of neutrons?
- Why do isotopes of same elements have same chemical but different physical properties?
- How the isotopes of Carbon are different from isotopes of Hydrogen?

Summary

- The Electron is lightest particle carrying a negative charge in an Atom discovered by J.J.Thomson and William Crooks.
- The Proton is positively charged particle discovered by Goldstein in 1886.J.J.Thomson investigated properties of proton in 1897.
- In 1932 Chadwick become successful to discover Neutron.
- Lord Rutherford in 1911, carried out series of experiments and proposed a new model for the atom that an atom contains nucleus at the center and electrons revolve around this nucleus.
- In 1913 Neil Bohr proposed another atomic model. This atomic model was different in this manner that it shows two folds, first to remove the Rutherford atomic model and second explain the line spectrum of Hydrogen atom based on quantum theory of Max Planck.
- In 1923 Lois De Broglie extend the wave particle duality to electron, and propose a hypothesis that all matter has particle as well as wave nature at the sub microscopic level.
- The Energy levels or Shell or Orbit are all possible paths on which electrons are revolving around nucleus. Which is shown by 'n'. these Shells are named as K, L, M, N, O, P.
- Main energy level are divided in to sub energy levels and known as sub shells.
- The distribution of electrons among the different orbits/shells and subshells is known as the electronic configuration of an atom.
- Atoms of the same elements having same atomic number but different atomic masses are called isotopes. They have same number of electron and same number of protons, but different number of neutrons.
- The Isotopes are used in worldwide applications of daily life. Research laboratories, medical centers, industrial facilities, food irradiation plants and many consumer products all use or contain isotopes.

EXERCISE

SECTION- A: MULTIPLE CHOICE QUESTIONS

Tick Mark (✓) the correct answer

1. In an atom number of protons and neutrons are added to obtain:
   (a) number of electrons (b) number of nucleons
   (c) atomic number of element (d) number of isotopes
2. If proton number is 19, electron configuration will be:
   (a) 2, 8, 9 (b) 2, 8, 8, 1
   (c) 2, 8, 1 (d) 2, 8, 3
3. If nucleon number of potassium is 39 and its atomic number is 19 then, number of neutrons will be:
   (a) 39 (b) 19
   (c) 20 (d) 29
4. The isotope C-12 is present in abundance of:
   (a) 96.9% (b) 97.6%
   © 98.89% (d) 99.7%
5. Electronic configuration is distribution of:
   (a) protons (b) neutrons
   (c) electrons (d) positrons
6. Which one of the following is most penetrating?
   (a) electron (b) Proton
   (c) alpha particle (d) neutron
7. How many subshells in a L shell:
   (a) one (b) two
   (c) three (d) four
8. De Broglie extend the wave particle duality to electron in:
   (a) 1920 (b) 1922
   (c) 1923 (d) 1925
9. Name the material of screen which is used in Rutherford atomic model :
   (a) Aluminum foil (b) zinc sulphide
   (c) sodium sulphide (d) Aluminum sulphide
10. Which rays are used for sterilization of medical instruments :
   (a) α-rays (b) β-rays
   (c) γ-rays (d) x-rays

SECTION- B: SHORT QUESTIONS:

1. Draw the structure of isotopes of chlorine to justify the definition of isotopes?
2. An atom has 5 electrons in M shell than:
   (a) Find out its atomic number?
   (b) Write Electronic configuration of atom?
   (c) Name the element of atom?
3. Justify that Rutherford atomic model has defects?
4. Describe wave particle duality of electron of De Broglie Hypothesis?
5. What are Limitations of Bohr's Atomic Model?
6. Differentiate between shell and sub shell with examples?
7. How the atoms of O_8^{17} and O_8^{16} are similar or different from each other?
9. Write down the names of sub atomic particles their masses in a.m.u with their unit charges.

SECTION- C: DETAILED QUESTIONS:

1. Discuss Rutherford' gold metal foil experiment in the light of structure of atom.
2. Writ down the applications of isotopes in daily life.
3. Explain how Bohr's atomic model is different from Rutherford atomic model.
4. Prove that modern theory of De Broglie is related with Einstein and Plank's equations.
5. How are cathode rays produced? What are their major characteristics?
6. Describe the schrodinger atomic model.
7. Describe briefly the experiments which provide clue and evidences of electron, proton and neutron in an atom.
8. How many protons, neutrons and electrons are present in the following elements ?
   (i) Fe_{26}^{56} (ii) O_8^{17} (iii) Cl_{17}^{37} (iv) U_{92}^{235} (v) C_6^{14}

Chapter 3

PERIODIC TABLE AND PERIODICITY OF PROPERTIES

Time Allocation

Teaching periods	=	08
Assessment period	=	02
Weightage	=	08

Major concepts

3.1 Periodic Table.

3.2 Periodicity of properties.

STUDENTS LEARNING OUT COMES (SLO'S)

Student will be able to:

- State the periodic law.
- Distinguish between a period and a group in the periodic table.
- Classify the elements (into two categories: groups and periods) according to the configuration of their outer most electrons.
- Determine the demarcation of the periodic table into an s-block, p-block, d-block and f-block
- Construct the shape of the periodic table.
- Determine the location of families on the Periodic Table.
- Recognize the similarity in the chemical and physical properties of elements in the same family of elements.
- Identify the relationship between electron configuration and the position of an element on the periodic table.
- Explain how shielding effect influences periodic trends.
- Describe how electro negativitiy, electron affinity, atomic radii and ionization energy change within a group and within a period in the periodic table.

Introduction:

The periodic table of Elements which you see in front of any classroom or chemistry laboratory. You take it for granted but this is the product of hundreds years struggle of scientists to understand the complexity of this world as Elements . When a large number of elements discovered, scientists decided to arrange the elements in certain order.

First of all German chemist Dobereiner proposed classification of Triads in which several groups of three elements classified on the basis of atomic masses. In this Triad central element had atomic mass approximately equal to average of mass of the other two elements. For example, Calcium (40), Strontium (87.6) and barium (137). In which mass of strontium is average of atomic masses of calcium and barium.

Table 3.1 Dobereiner classification of traids

ELEMENTS		ATOMIC MASS	ARITHMATIC MEAN
Traids	Lithium Sodium Potassium	7 23 39	$\frac{7 + 39}{2} = 23$
Triads	Chlorine bromine Iodine	35.5 80 126.5	$\frac{35.5 + 126.5}{2} = 81$
Triads	Calcium Strontium Barium	40 87.6 137	$\frac{137 + 40}{2} = 88.5$

In 1864 British chemist Newland put forwarded Law of Octaves and arranged the elements in order of increasing atomic masses. According to him if elements are arranged in increasing order of their masses then eighth element has similar properties as of first element in a group. For example:

Table 3.2 Newlands Classification octaves

Li=7	Be=9	B=11	C=12	N=14	O=16	F=19
Na=23	Mg=24	Al=27.3	Si=28	P=30	S=32	Cl=35.5

In the above arragement Li and Na, Be and Mg, B and Al, C and Si, N and P, O and S, F and Cl shows same chemical properties.

In 1869 Mendeleev published a periodic table containing eight vertical columns (groups) and horizontal rows(periods)on the basis of physical and chemical properties of elements. In 1869 German scientist Lother Meyer published a periodic table in which 56 elements were arranged in 9 vertical columns or groups on the basis of atomic masses


Dobereiner

Newland

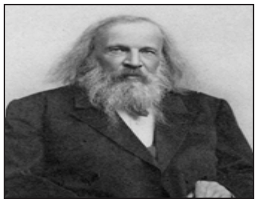
Lother Meyer

Mendeleev

Fig 3.1 Scientists participated in classification of Periodic table

3.1 PERIODIC TABLE

Mendeleev's periodic table was the first attempt to arrange the elements although this periodic table was failed due to many demerits but provided the base for discovery of Periodic Law. On the basis of periodic Law a periodic table developed in which vertical columns are called groups and horizontal rows are called periods. This periodic table predict the properties of elements.

3.1.1 Periodic Law:

In 1869 Mendeleev Proposed a periodic law on the basis of physical and chemical properties empirically. Periodic law states that " The Properties of the elements are a periodic Function of their atomic weight". In certain cases Mendeleev left gaps, which were modified by Moseley.

3.1.2 Modern Periodic Table

Atomic number is fundamental property because it increases regularly element to element and is fixed for every element. It was noticed in arrangement of elements that atomic number increasing from left to right in a horizontal row and properties of elements were found repeating after regular intervals. Due to this reason elements of same properties and same electronic configuration are placed in the same group.

In 1913 Moseley discovered that Atomic number is the basic property of an atom. He proposed a modern periodic law. The Moseley states that " The Physical and chemical Properties of elements are the periodic function of their atomic numbers" Atomic number of an element is equal to the number of electrons in a neutral atom so atomic number also provides the electronic configuration of elements of periodic table. So on the basis of Electronic configuration elements are arranged in long form of periodic table and Periodic table is composed with 7 rows and 18 columns.

Periods in Periodic Table:

There are seven horizontal rows in periodic table known as periods. In periods physical and chemical properties changes from left to right. Elements of a period shows different properties because The electronic configuration continuously changing within a period and number of valence electrons decide the position of element in a period. These periods are categorized as short periods and long periods. Which are as follows.

First Period(shortest period)

- This period contains only two elements Hydrogen (H) and Helium (He).
- K-shell is filled in this period.

Second and Third Period(Short Period)

- Each period contain eight elements.
- In these Periods L and M shells are being filled by electrons.
- Second period contains Li, Be, B, C, N, O,F and Ne.
- Third period contains Na, Mg, Al, Si, P,S, Cl, and Ar.

Fourth and Fifth Period(Long Period)

- Each period contain 18 elements.
- In these periods M and N shells are being filled by electrons.
- Fourth period starts from Potassium (K) and ends on Krypton (Kr).
- Fifth period starts from Rubidium (Rb) and ends on Xenon (Xe).

Sixth Period(Longest Period)

- This period contains 32 elements.
- The 14 elements in the bottom are named as Lanthanides.
- Sixth periods starts from Caesium (Cs) and ends with Radon (Rn).

Seventh Period (Incomplete Period)

- This period starts from Francium (Fr)
- This period is consider as incomplete.
- This period contains a group of 14 elements known as Actinides.

All the periods except the first period start with an alkali metal and end at a Nobel gas. It is observed that number of elements are fixed in each period because of maximum number of electrons accommodation in the particular valence shell of elements. Which is shown in the following table 3.2.

Table 3.2 Period wise Atomic Number of Elements in Periodic Table

Period Number	Number of Elements	Range of Atomic Number
First	2	1 to 2
Second	8	3 to 10
Third	8	11 to 18
Fourth	18	19 to 36
Fifth	18	37 to 54
Sixth	32	55 to 86
Seventh	[32]*	87 to 118*

(Where * Shows incomplete period)

Periodic Table of elements

Atomic Number
Symbol
Name
Atomic Mass

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 1.008																	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.086	15 P Phosphorus 30.974	16 S Sulfur 32.066	17 Cl Chlorine 35.453	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.631	33 As Arsenic 74.922	34 Se Selenium 78.971	35 Br Bromine 79.904	36 Kr Krypton 84.798
37 Rb Rubidium 84.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.95	43 Tc Technetium 98.907	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.906	46 Pd Palladium 106.42	47 Ag Silver 107.868	48 Cd Cadmium 112.411	49 In Indium 114.818	50 Sn Tin 118.711	51 Sb Antimony 121.760	52 Te Tellurium 127.6	53 I Iodine 126.904	54 Xe Xenon 131.294
55 Cs Cesium 132.905	56 Ba Barium 137.328	57-71	72 Hf Hafnium 178.49	73 Ta Tantalum 180.948	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.217	78 Pt Platinum 195.085	79 Au Gold 196.967	80 Hg Mercury 200.592	81 Tl Thallium 204.383	82 Pb Lead 207.2	83 Bi Bismuth 208.980	84 Po Polonium [208.982]	85 At Astatine 209.987	86 Rn Radon 222.018
87 Fr Francium 223.020	88 Ra Radium 226.025	89-103	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [269]	109 Mt Meitnerium [268]	110 Ds Darmstadtium [269]	111 Rg Roentgenium [272]	112 Cn Copernicium [277]	113 Uut Ununtrium unknown	114 Fl Flerovium [289]	115 Uup Ununpentium unknown	116 Lv Livermorium [298]	117 Uus Ununseptium unknown	118 Uuo Ununoctium unknown

Lanthanide Series	57 La Lanthanum 138.905	58 Ce Cerium 140.116	59 Pr Praseodymium 140.908	60 Nd Neodymium 144.243	61 Pm Promethium 144.913	62 Sm Samarium 150.36	63 Eu Europium 151.964	64 Gd Gadolinium 157.25	65 Tb Terbium 158.925	66 Dy Dysprosium 162.500	67 Ho Holmium 164.930	68 Er Erbium 167.259	69 Tm Thulium 168.934	70 Yb Ytterbium 173.055	71 Lu Lutetium 174.967
Actinide Series	89 Ac Actinium 227.028	90 Th Thorium 232.038	91 Pa Protactinium 231.036	92 U Uranium 238.029	93 Np Neptunium 237.048	94 Pu Plutonium 244.064	95 Am Americium 243.061	96 Cm Curium 247.070	97 Bk Berkelium 247.070	98 Cf Californium 251.080	99 Es Einsteinium [254]	100 Fm Fermium 257.095	101 Md Mendelevium 258.1	102 No Nobelium 259.101	103 Lr Lawrencium [262]

Alkali Metal	Alkaline Earth	Transition Metal	Basic Metal	Semimetal	Nonmetal	Halogen	Noble Gas	Lanthanide	Actinide

Groups in Periodic Table:

There are Eighteen vertical columns in periodic table known as groups. The sub groups are divided on the basis of their similar properties as A and B and placed together in periodic table.

The elements of sub group A are called Main or Representative Elements.

The elements of sub group B are called Transition Elements. The group number indicate total number of electrons in valence shell of the element.

Group I A (Alkali Metal) or Lithium Family:

- This Group include Lithium (Li), Sodium (Na), Potassium (K), Rubidium (Rb), Cesium (Cs) and Francium (Fr).
- Their Valence shell contain one electron.
- On reaction they lose one electron and form univalent positive ion (cation).
- They are highly reactive metals.
- They have low melting point.

Do you know?

Francium(Fr) is radioactive element of IA group.

Group II A(Alkaline Earth Metals) or Beryllium Family:

- This Group include Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) and Radium (Ra).
- Their Valence shell contain two electrons.
- On reaction they lose two electrons and form divalent positive ion.
- They show irregular Densities, Melting and Boiling point.

Do you know?

Radium (Ra) is radioactive element of IIA group.

Group III A(Boron Family) :

- This Group include Boron (B), Aluminum (Al), Gallium (Ga), Indium (In) And Thallium (Tl).
- Their valence shell contain three electrons.
- On reaction they lose three electrons and form trivalent positive ion except Boron.

Do you know?

Boron (B) is metalloid in III A group due to increase in atomic volume boron shows some properties of metals and some properties of non-metals.

Group IV A(Carbon Family) :

- This Group include carbon (C),silicon (Si), Germanium (Ge), Tin (Sn) and Lead (Pb).
- Their valence shell contain four electrons.

- C,Si and Ge form covalent bond, whereas Sn and Pb exhibit variable Valencies 2 and 4.
- Carbon is nonmetal, Silicon , Germanium are metalloids and Tin and Lead are metals.

Group V A(Nitrogen Family) :

- This Group include Nitrogen (N), Phosphorus (P), Arsenic (As), Antimony (Sb) and Bismuth (Bi).
- Their valence shell contains five electrons.
- They show large variations in their properties as we go down the group.
- Except Nitrogen all exist in allotropic form.

Do you know?

Carbon and Tin exist in allotropic form in IVA group, due to increase in atomic radii and volume, addition of new shell takes place.

Group VI A(Oxygen Family) :

- This Group include Oxygen(O) ,Sulphur (S),Selenium (Se), Tellurium (Te) and Polonium (Po).
- Their valance shell contain six elements.
- All of these elements exist in allotropic forms.
- Oxygen and sulphur are nonmetals, polonium is metal and all other are metalloids.

Group VII A(Halogen Family) :

- This Group include Fluorine (F),Chlorine(Cl),Bromine(Br),Iodine(I) and Astatine(At).
- Their valence shell contain seven electrons.
- Except Astatine (metal) all are nonmetals.
- Fluorine and chlorine are gases, bromine is liquid and iodine is solid at room temperature.

Group VIII A(Inert or Nobel gases):

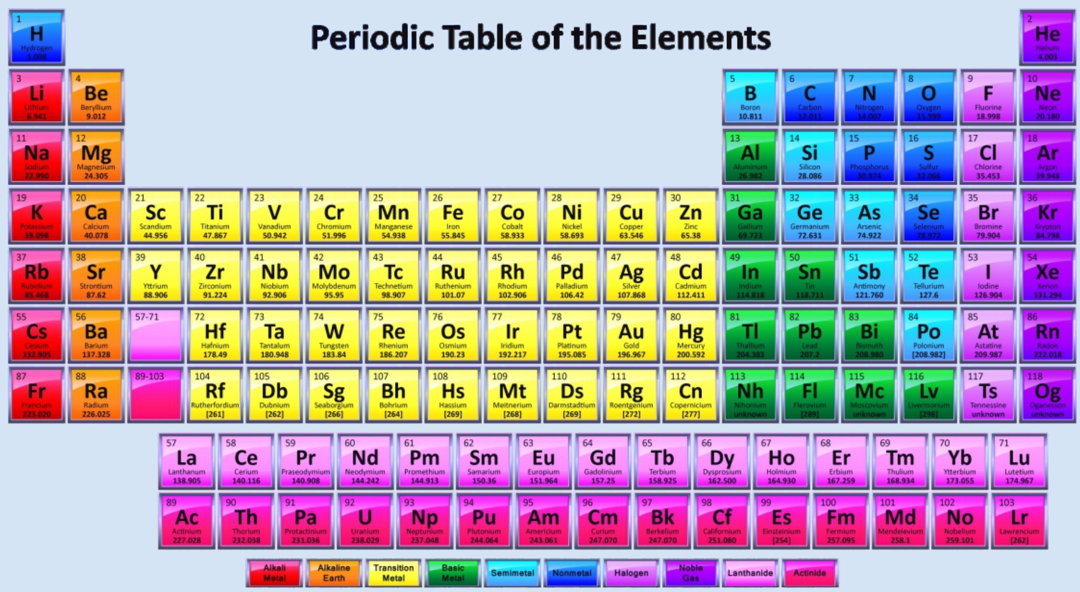
- This Group include Helium(He), Neon(Ne), Argon(Ar), Krypton (Kr) Xenon(Xe) and Radon(Rn).
- Their valence shell contain eight electrons except Helium which contain two electrons.

Group IB to VIII B(Transition Elements):

- These Groups are metals.
- In chemical reactions they shows Variable valencies.
- Their valence shells are incomplete.

Test Yourself

(1) look at the given periodic table carefully and answer the following questions



- Identify and list down the solid, liquid and gases at room temperature from the given periodic table.
- Identify and name the artificial elements from the periodic table gives above.
- Identify and list down the radioactive elements.
- Identify alkali , alkaline , transition metals.
- Identify and list down metalloids , lanthanide and actinide.

3.1.3 Demarcation of periodic table in s, p, d and f blocks:

The periodic table has been divided into four blocks, s , p , d , and f based on electronic configuration.

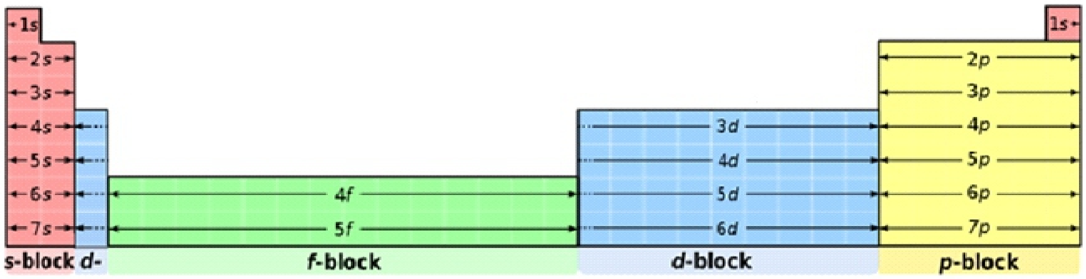


Fig 3.2

Nobel Gases:They are colorless, unreactive and diamagnetic,They are placed in zero group. Their electronic configuration is ns^2 ,np^6 and are exceptionally stable.

Representative Elements It includes metals and nonmetals. Some are diamagnetic and some are paramagnetic and marked as S block and P block elements.

(I) s-Block Elements: In **s** block elements electrons occupy in ns orbital. The elements of group IA and IIA are s block elements. Their electronic configuration varies ns^1 to ns^2.

(II) p-Block Elements: In **p** block elements electrons begin to fill np^1 to np^6. Elements of group IIIA to VIIA and zero group except He are also p block elements.

d-block Elements (Outer Transition Elements): These elements are metals but their properties are different from metals of representative elements i.e. Melting point, Boiling Point, variable oxidation state color compounds etc. In these elements electron fills in ns^2 $(n-1)$ $d^{1\text{-}10}$ orbital. d-block elements consist of three series.

f-Block Elements (Inner Transition Elements): The elements in which inner f-orbital is filled, are called f block elements. They exhibit electronic configuration $ns^2(n-1)d^1(n-2)f^{1-14}$ There are two series called Lanthanides and Actinides.

3.2 PERIODICITY OF PROPERTIES:

The Periodicity means" Repeatation of some thing after some interval". The Periodicity of properties means that elements are arranged in an Order where properties of elements repeat after some period.

3.2.1 Atomic Size and Atomic Radius:

Atoms are so small that it is impossible to see atoms even with a powerful optical microscope. The size of a single atom therefore cannot be directly measured. however, techniques have been developed which can measure the distance between the centres of two bonded atoms of any elements. Half of this distance is considered to be the radius of the atom. It is measured in Angstrom unit (A) 1A = 1 0^{-8} cm

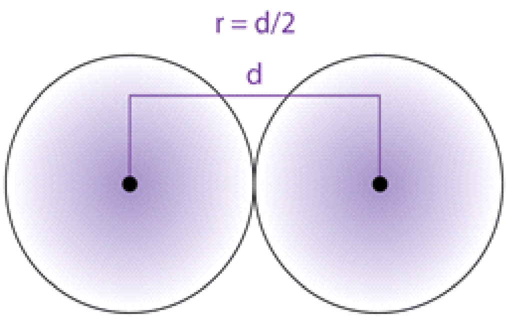


Fig 3.3 Atomic radius

In the periodic table, that atomic radius increases from top to bottom within a group due to increase in number of shells. However, as the atomic number increases from left to right, the atomic radius decrease. This gradual decrease in the radius is due to increase in the positive charge on the nucleus. As the positive nuclear charge increases, the negatively charged electrons in the shells are pulled closer to the nucleus. Thus, the size of the outermost shell becomes gradually smaller. This effect is quite remarkable in the elements of longer periods in which " d " and " f " sub shells are involved. For example, the gradual decrease in the size of Lanthanides is significant and called Lanthanide Contraction.

Table 3.4 Atomic radii decreases in period

2nd Periods elements	^{3}Li	^{4}Be	^{5}B	^{6}C	^{7}N	^{8}O	^{9}F	^{10}Ne
Atomic radii (pm)	152	113	88	77	75	73	71	69

Table 3.5 Atomic radii increases in group

Note:
1A° = 100pm

1st group elements	Atomic radii (pm)
^{3}Li	152
^{11}Na	186
^{19}K	227
^{37}Rb	248
^{55}Cs	265

3.2.2 Ionization Energy:

The ionization energy is minimum amount of energy required to remove an electron from a gaseous state and measured in joule/mole. The ionization energy depends upon atomic size and nuclear charge. The higher ionization energy means removal of electron is more difficult for example the ionization energy of hydrogen is 1312KJ/mole.

$$H(g) + \text{Energy} \longrightarrow H^{+}_{(g)} + e^{-} \quad (\text{I.E} = +1312\text{KJ/mol})$$

If we move from left to right in periods the value of ionization energy increases. Its because of size of atoms reduces and electrons are held strongly by the attractive force of nucleus. Due to this elements on the left side have less ionization energy. Which is shown in table 3.6.

Table 3.6 Ionization energy increases in period

Increase in ionization Energy of Elements in Periods of Periodic Table

Elements of Second periods	^{3}Li	^{4}Be	^{5}B	^{6}C	^{7}N	^{8}O	^{9}F	^{10}Ne
Ionization Energy(kJ/mol	520	899	801	1086	1402	1314	1681	2081

As we move down the group ionization energy decreases from top to bottom due to additions of shells. Decrease in ionization energy is shown in table 3.7.

Increase in number of shells reduce the electrostatic force between electrons of valence shell and nucleus.

Table 3.7 Ionization Energy Decreases in group

Elements of first group	Ionization energy (KJ/mol)
^{3}Li	520
^{11}Na	496
^{19}K	419
^{37}Rb	403
^{55}Cs	377

3.2.3 Electron Affinity

The electron affinity is amount of energy released when an electron is added in the outermost shell of a gaseous atom. It is also calculated in K J/mol. Affinity means attraction, therefore electron affinity means tendency to accept electron to form anion For example electron affinity of fluorine is -328 KJ/mol.

$$F_{(g)} + e^{-} \longrightarrow F^{-}_{(g)} \quad \Delta H = -328 \text{ KJ/mol}$$

In a period electron affinity increases from left to right due to decrease of atomic size because when size of atom decreases the attraction between nucleus and incoming electron increases and more energy is released.

Table 3.8 Electron affinity increases in period

Elements of Second periods	^{3}Li	^{4}Be	^{5}B	^{6}C	^{7}N	^{8}O	^{9}F	^{10}Ne
Electron Affinity (kJ/mol	-60	-48	-29	-122	-6.8	-141	-328	0

Electron Affinity decreases in group

In a group electron affinity values decrease from Top to bottom, because the size of atom increases.

Table 3.9 Electron Affinity decreases in Groups

Elements of 17th group	Electron Affinity (KJ/mol)
^{9}F	-328
^{17}Cl	-349
^{35}Br	-325
^{53}I	-295

Down the group attraction of incoming electron and Nucleus decreases and less energy released.

As the size of iodine is bigger than bromine its electron affinity is less than bromine.

The decrease of electron affinity is shown in table 3.9

3.2.4 Shielding Effect:

The shielding effect is defined as reduction in the effect of nuclear charge on the valence electron cloud, due to presence of inner shell electrons in an atom.

The electrons present between the nucleus and valence shell of atom reduce the nuclear charge effect on electrons present in outermost shell. As a result valence electron experience less nuclear charge than the actual charge. Therefore

"Electrons present in the inner shells Shield the force of attraction of nucleus felt by the valence shell electrons is called Shielding effect."

The Shielding effect increases down the group in periodic table and remain same in period from left to right. For example shielding effect in potassium atom is more than sodium atom.

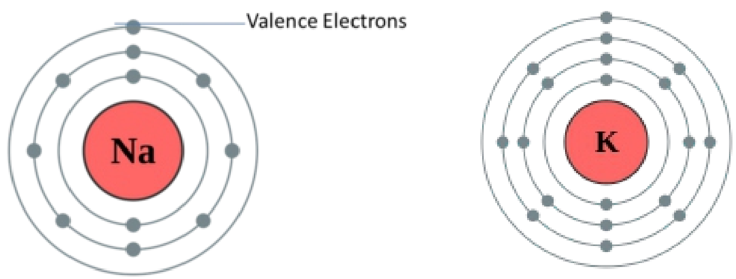


Fig 3.4 Shielding effect in Potassium atom is more then Sodium atom

3.2.5 Electronegativity:

The ability of an atom to attract the shared pair of electrons towards itself in a molecule is called electronegativity. The trend of electronegativity is same as ionization energy and electron affinity. It increases from left to right in period due to increase in nuclear charge which decrease the distance from nucleus to shared electron pair (table 3.10). It increases the power to attract the shared pair of electrons.

Table 3.10 Electronegavity Increases in periods

Elements of Second periods	^{3}Li	^{4}Be	^{5}B	^{6}C	^{7}N	^{8}O	^{9}F
Electronegativity	1.0	1.6	2.0	2.6	3.0	3.4	4.0

In group electronegativity decreases because size of Atom increases and attraction for shared electron pair decreases: for example in table 3.11 Electronegativity of halogens are given.

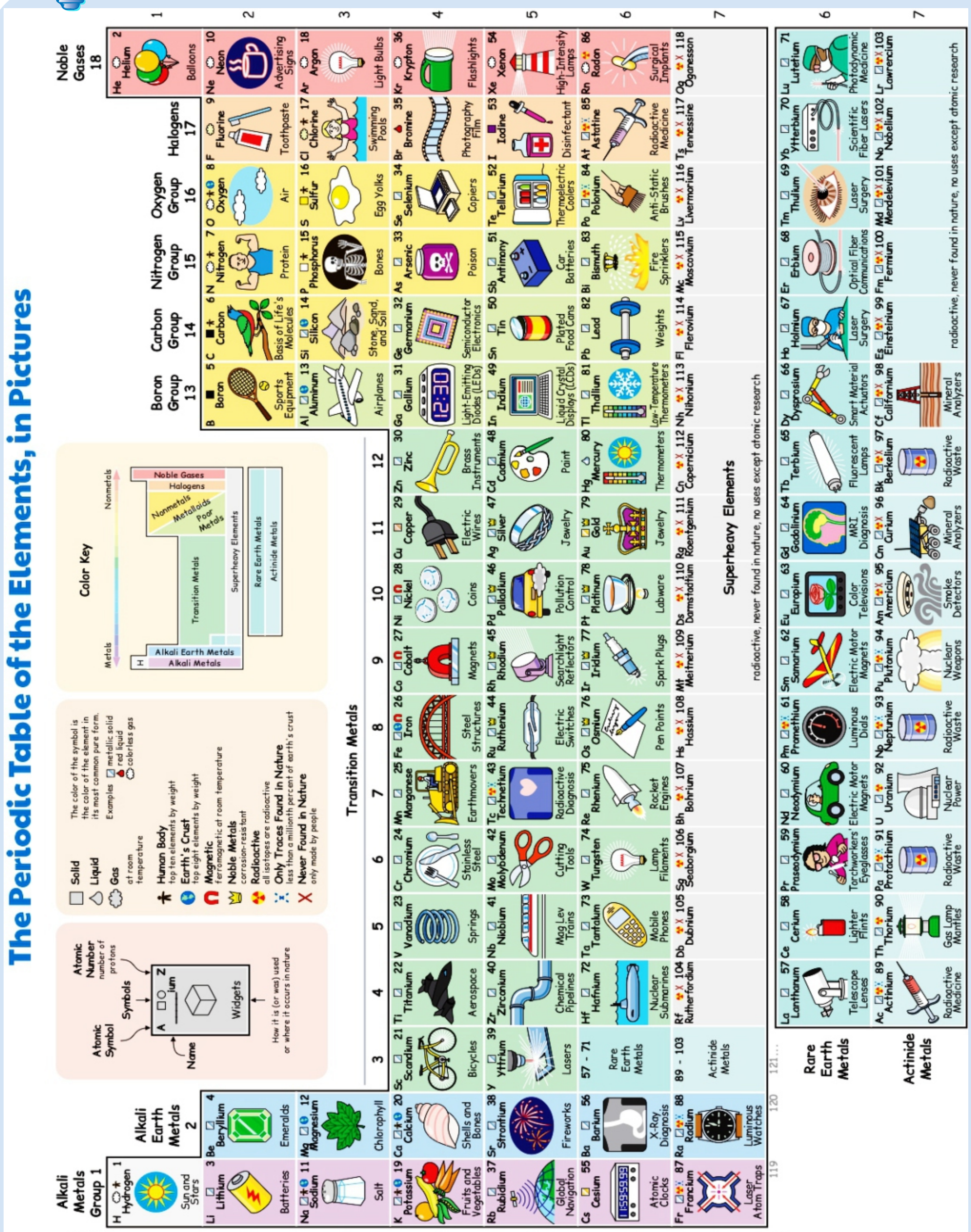
Table3.11 Electronegativity decreases in group

Elements of 17th group	Electronegativity
^{9}F	4.0
^{17}Cl	3.2
^{35}Br	3.0
^{53}I	2.7

Test Yourself

- What is trend of atomic radius in group?
- Why do bigger size atoms have more shielding effects?
- Which element have highest ionization energy and why?

Do you know?


The Periodic Table of the Elements, in Pictures
Color Key
Transition Metals
Superheavy Elements
Rare Earth Metals
Actinide Metals
Alkali Metals Group 1
Alkali Earth Metals 2
Boron Group 13
Carbon Group 14
Nitrogen Group 15
Oxygen Group 16
Halogens 17
Noble Gases 18
Periods

Summary

- 19th century is considered as milestone in Systematic arrangment of elements in Periodic Table.
- Dobereiner arranged elements in Triads.
- Newland put forward Law of Octaves.
- Mendeleev Published Preiodic law with groups and Rows.
- Moseley stated his law as "The physical and chemical properties of elements are the periodic function of their atomic numbers"
- There are total eighteen groups and seven periods in modern periodic table.
- Physical and Chemical properties change from left to right in a period. Elements of a period show different properties because the electronic configuration continuously changes within a period.
- The sub groups are divided on the basis of their similar properties as A and B and placed together in periodic table.
- The elements of sub group A are called Main or Representative Elements.
- The elements of sub group B are called Transition Elements. The group number indicate total number of electrons in valence shell of the element.
- The atomic size increases down a group but decreases along the period.
- Ionization energy decreases down a group but increases along a period.
- Electronegativity decreases down a group but increases along a period.
- Electron affinity decreases down a group but increases along a period.
- The Shielding effect increases down the group in periodic table and remain same in period from left to right.

EXERCISE

SECTION- A: MULTIPLE CHOICE QUESTIONS

Tick Mark (✓) the correct answer

1. In 1869 Mandeleev put forward his periodic law about:

   (a) Atomic Number (b) Chemical properties

   (c) Physical properties (d) Atomic Mass

2. The periodic table divided into s, p, d, and f block based on.

   (a) Atomic Radius (b) Electronic Configuration

   (c) Ionization Energy (d) Electron Affinity

3. 4th and 5th period in periodic table are known as:

(a) Short period (b) Long period

(c) Normal period (d) Very long period

4. Which one of the following decreases along the period?

(a) Ionization Energy (b) Atomic Radius

(c) Electronegativity (d) Electron Affinity

5. The elements of VIIA group are known as:

(a) Lanthanides (b) Actinides

(c) Halogens (d) Nobel Gases

6. According to Mosely the chemical properties of elements are the periodic function of their :

(a) Atomic Size (b) Atomic Mass

(c) Atomic Radius (d) Atomic Number

7. The shielding effect across the period :

(a) Increases (b) Decrease

(c) Moderate (d) Same

8. The ability to attract shared pair of electron is called:

(a) Electron Affinity (b) Electronegativity

(c) Ionization Energy (d) Shielding Effect

9. In group electron affinity values decreases from top to bottom because :

(a) Atomic size normal (b) Atomic size increases

(c) Atomic size decreases (d) Atomic size same

10. All Transition Elements are :

(a) Gases (b) Metals

(c) Nonmetals (d) Metalloids

SECTION- B: SHORT QUESTIONS:

1. Distinguish between periods and groups.
2. Describe the trend of electronegativity within group and period with the help of examples?
3. Explain the similarity of chemical and physical properties of elements in the same family.
4. Justify that periodicity of properties dependent upon number of protons in an atom?
5. Identify that which halogens exist as gases, liquid and solid?
6. Why Alkaline earth metals shows irregular melting and boiling point?
7. Why ionization energy, electron affinity and electronegativity exhibit same trend in period and groups?

SECTION- C: DETAILED QUESTIONS:

1. Discuss in detail the long form of periodic table.
2. Determine the demarcation of periodic table in to s, p, d and f blocks.
3. Identify the electronic configuration of the following elements.

   Na, Ca, F, Si
4. Determine the location of families on periodic table.
5. Discuss that Mendlevee periodic law provide a base for modern periodic table.
6. Explain how shielding effect influence the periodic trends?

Chapter 4

CHEMICAL BONDING

Time Allocation

Teaching periods	= 14
Assessment period	= 4
Weightage	= 14

Major Concepts

4.1 Why do atoms form chemical Bonds?
4.2 Formation of Chemical Bonds.
4.3 Types of Bonds.
4.4 Intermolecular Forces.
4.5 Nature of Bonding and Properties.

STUDENTS LEARNING OUTCOMES (SLO'S)

Students will be able to:

- Find the number of valence electrons in an atom using the Periodic Table.
- Describe the importance of noble gas electronic configurations.
- State the octet and duplet rules.
- Explain how elements attain stability.
- Describe the ways in which bonds may be formed.
- State the importance of noble gas electronic configurations in the formation of ion.
- Describe the formation of cations from an atom of a metallic element.
- Describe the formation of anions from an atom of a non-metallic element.
- Describe the characteristics of an ionic bond.
- Recognize a compound as having ionic bonds.
- Identify characteristics of ionic compounds.
- Describe the formation of a covalent bond between two non metallic elements.
- Describe with examples single, double and triple covalent bonds.
- Describe the properties of polar and non-polar compounds.
- Draw electron cross and dot structures for simple covalent molecules containing single, double and triple covalent bonds.
- Describe the weak forces of interaction such as dipole-dipole interaction and hydrogen bonding.

INTRODUCTION:

In the previous chapter, you have learnt about the matter. You are also familiar with the fact that all the matters in this world are composed of atoms. The attractive force which binds atoms together is called as a chemical bond or chemical forces. Few elements also consist of un-bounded atoms. For instance, helium, neon, argon, xenon and krypton present in the atmosphere consist of un-bounded atoms. The way in which various atoms are bonded together has a effect on the properties of substances.

In this chapter, we will explore the nature of various types of chemical bonding.

4.1 WHY DO ATOMS FORM CHEMICAL BOND?

Why do atoms form chemical bonds? The essential answer is that everyone in the world desires to be stable in their life. Atoms are just like that, they are also trying to become more stable, so atom try to shares required number of electrons with each other to obtain the electronic configurations of noble gases.

Electronic Configuration of Noble Gas

Noble gases have $ns^{2,}$ np^6 electronic configuration in the outermost shell and rarely form chemical bonds. The noble gases are Helium (He), Neon (Ne), Argon (Ar), Krypton (Kr), Xenon (Xe) and Radon (Rn). These elements are sometimes called the inert gases. This is because they do not participate in chemical reactions. Outer shell of five noble gas atoms are shown in figure 4.1

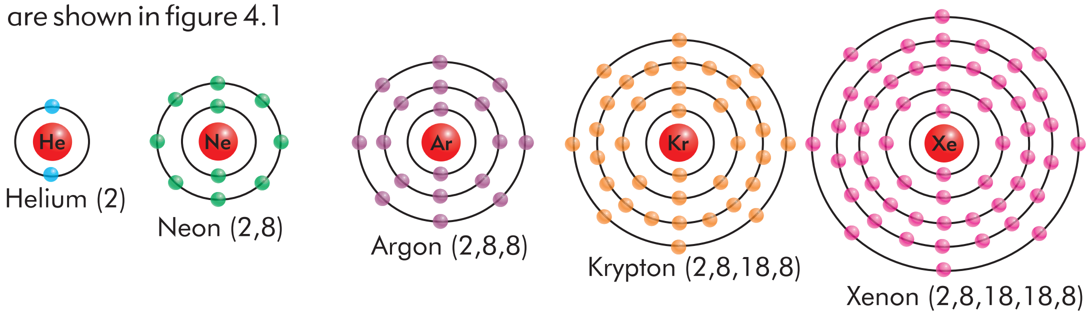


Figure: 4.1 Outer electronic configuration of the noble gases

Note that Helium contains 2 electrons and other noble gases contain 8 electrons in the valence shell. Because of these electronic configurations noble gases are stable and not active. Atoms to acquire two electrons in the valence shell is called duplet rule, whereas atoms to acquire eight electrons in the valence shell is called the octet rule. In 1916 a chemist G.N. Lewis used this fact, why atoms undergo chemical reactions. He called his explanation as octet rule. An octet means a set of eight.

What are valence electrons?

Electrons present in the outermost shell of any atom play an important role in determining the chemical properties of the atom, including its ability to form chemical bonds. These electrons in the outermost shell of an atom are called as valence electrons or outer electrons. Finding of valence electron or the electron configuration consider an example of Boron (B), it has electronic number five. The electronic configuration looks like this: $1s^2$, $2s^2$, $2p^1$ since there are three electrons in the second shell ($2s^2$and $2p^1$), we can say boron has three valence electrons. The valence electrons which are involved in chemical bonding are termed as bonding electrons.

In the chapter 3, you have learned that the group number indicates the number of valence electrons in an atom. For example, sodium belongs to group IA, so it contains one electron in its valence shell. Similarly, phosphorus belongs to group VA, so it contains five electrons in the valence shell.

Test Yourself

- Why an atom form chemical bond?
- When atoms are considered to be unstable?
- Why doesn't helium atom tend to gain electron?
- Where are valence electrons located, and why are they important?
- What is meant by bonding electrons?
- Write the electronic configuration of Ne (atomic No. 10),Carbon (atomic No. 6) and sulphur (atomic No. 16).
- Why, noble gases do not react with other element to form compounds?
- Find the number of valence electrons in the following atoms.
  (a) Chlorine (b) Sodium
  (c) Magnesium (d) Potassium

4.2 FORMATION OF CHEMICAL BOND

Chemical bonding is the combining of atoms to form new substances. An interaction that holds two atoms together is called a chemical bond. Atoms can lose, gain or share valence electrons to form chemical bonds.

4.3 TYPES OF CHEMICAL BONDS

There are three types of bonds depending on the tendency of an atom to lose or gain or share electrons.

1. Ionic Bond
2. Covalent Bond
3. Co-ordinate covalent bond or dative covalent

4.3.1 Ionic Bonds

In the formation of ionic bond, an atom loses electrons and changes into positive ion (cation). Whereas another atom gains electron and changes into negative ion (anion). These cations and anions have opposite charges. They attract one another by electrostatic force of attraction. The electrostatic force of attraction that holds the oppositely charged ions together is called as ionic bond or electrovalent bond.

Generally, ionic bond is formed between the atoms of two different groups, metal and non-metal. Compounds that contain ionic bonds are called ionic compounds. Such as sodium chloride, potassium chloride, magnesium fluoride etc.
Formation of ionic bond is explained in following examples.

Example 1:

The Reaction between Sodium and Chlorine

Sodium atom is a metal of IA group of the periodic table and has only one electron in the outer most shell. The electron arrangement of sodium atom is 2, 8, 1. By losing one electron from the outer most shell, sodium forms cation (Na^+). Whereas chlorine atom is non-metal of VIIA group and has seven electrons in its outermost shell. The electron arrangement of chlorine atom is 2, 8, 7. Since chlorine atom has seven electrons in its outermost shell, it needs one electron to complete octet. By gaining one electron, chlorine atom now has eight electrons in its outermost shell and a chloride ion is formed (Cl^-).

$$Na \rightarrow Na^+ + e^-$$
$$2, 8, 1 \qquad 2, 8$$
$$Cl + \bar{e} \rightarrow Cl^-$$
$$2, 8, 7 \qquad 2, 8, 8$$

(Sodium chloride)

Both these atoms are now oppositely charged ions. Therefore two charged ions are attracted to each other by electrostatic force of attraction. Thus Na^+ and Cl^- ions are joint by ionic bond and form sodium chloride. The formation of ionic bonds by a 'dot and cross' diagram is shown in figure 4.2.

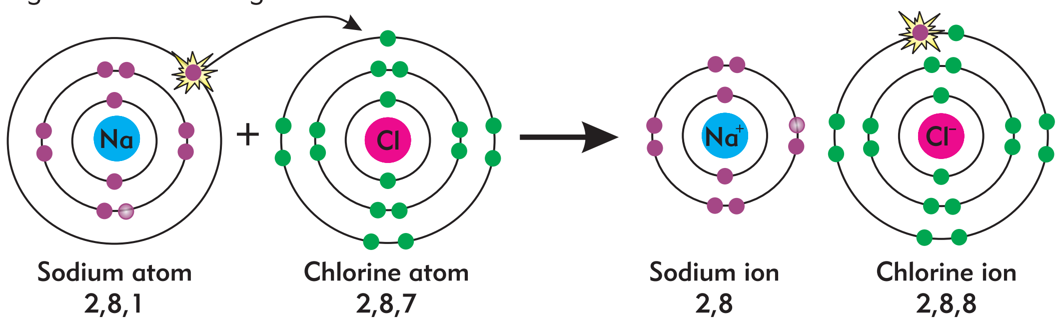


Fig 4.2 Formation of ionic bond in Sodium chloride

Example 2:

The Reaction between Magnesium and Oxygen

Consider another example of ionic bond formation is the reaction between magnesium and oxygen forming magnesium oxide. Magnesium is in IIA group of the periodic table and has only two electrons to share and oxygen is in group VIA and has six electrons in its outermost shell. By losing two electrons from the outermost shell, magnesium becomes Mg^{2+} and it is left with 8 electrons in the second shell. By gaining two electrons, oxygen atom now also has eight electrons in its outermost shell and becomes O^{2-}.Both these atoms are now changed into oppositely charged ions. The attraction between the oppositely charged ions forms the ionic bond between magnesium and oxygen. The formula of magnesium oxide is MgO. The formation of ionic bonds by a 'dot and cross' diagram is shown in figure 4.3.

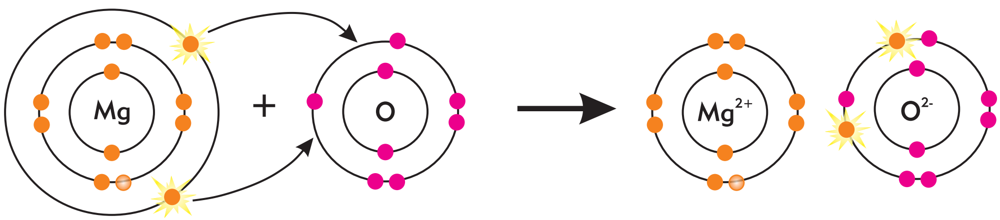


Fig 4.3 Formation of ionic bond in Magnesium oxide

The ionic bond between magnesium and oxygen is stronger than the ionic bond between sodium and Chlorine because of the greater charge density on the ions. Magnesium oxide has a higher melting point due to the presence of the stronger bond.

Do You Know?

- The alkali metals (IA group elements) lose a single electron to form a monovalent cation.
- The alkaline earth metals (IIA group elements) lose two electrons to form di valent cation (M^{++}).
- Aluminum, a member of the IIIA family, loses three electrons and form tri valent cation (M^{+++}).
- The halogens (VIIA group elements) have seven valence electrons. All the halogens gain one electron to complete their octet. And all of them form an anion with a single negative charge.
- The VIA elements gain two electrons and form di valent anions (eg O^{2-}, S^{2-})
- The VA elements gain three electrons and form tri valent anions (eg N^{3-}, P^{3-})

4.3.2 Covalent Bond

In this type of bond, electrons are not gained or lost by atoms. A covalent bond is formed by mutual sharing of electrons between two atoms. This type of bonding occurs between two atoms of the same element or atoms of different elements. This bonding occurs primarily between nonmetals; however, it can also be observed between metals and non-metals.

Consider the formation of covalent bond between two hydrogen atoms. Hydrogen has one electron in their valence shell. When two hydrogen atoms share their valence electrons, both atoms achieve the electronic configuration of noble gas and satisfy the duplet rule.

A covalent bond is generally represented by a short straight line (–) between two bonded atoms. Figure 4.4 shows the formation of a covalent bond by a 'dot' and 'cross' diagram.

Do You Know?

When there is one electron in a sub-orbital, it is called an unpaired electron. When the sub-orbital is filled with a maximum of two electrons, it is called an electron pair. The electron pairs can be found in two types as bond pair and lone pair.

The main difference between bond pair and lone pair is that bond pair is composed of two electrons that are in a bond whereas lone pair is composed of two electrons that are not in a bond.

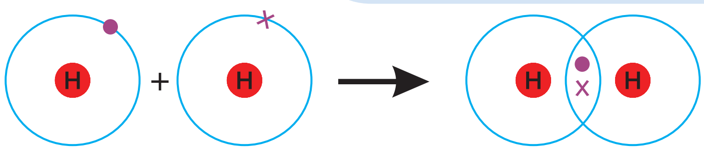


Fig 4.4 The formation of a hydrogen molecule

Types of Covalent Bond

As we know that the covalent bond is formed by mutual sharing of electrons between two atoms. The electrons of atoms that pair up to form a chemical bond are called bond pair electrons. Depending upon the number of bond pair, a covalent bond is further classified into three types.

- Single Covalent Bond
- Double Covalent Bond
- Triple Covalent Bond

Single Covalent Bond (–)

A covalent bond which is formed by the mutual sharing of one bond pair is called a single covalent bond and it is represented by a single short straight line. The formations of H-H, H-Cl, $CH_4$are few examples of this type of bonding. Below figure shows the formation of chlorine molecule by a dot and cross diagram.

Formation of Chlorine Molecule

A chlorine atom belongs to group VIIA and it has seven outer electrons. It needs one more electron to achieve a stable octet electronic configuration. When two chlorine atoms share their valence electrons, both atoms achieve the electronic configuration of noble gas. The single bond in chlorine molecule is represented by a dot and cross diagram as shown in figure 4.5.

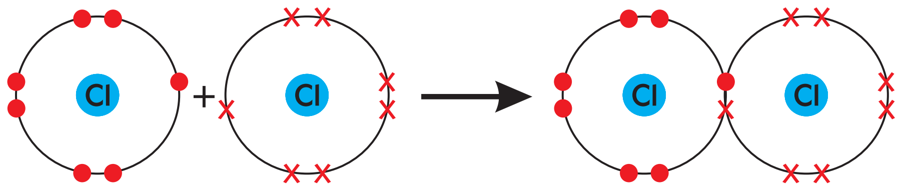


Fig 4.5 Formation of single covalent bond in chlorine molecule

Some other examples of formation of single covalent bond in hydrogen chloride and methane can be represented as follows:

H• + ∗Cl ⟶ H•∗Cl or HCl

Fig 4.6 Formation of single covalent bond in hydrogen chloride and methane

Double Covalent Bond (═)

A covalent bond which is formed by the mutual sharing of two bond pairs called a double covalent bond and it is represented by two short straight lines. The examples of molecules having double bonds are oxygen (O_2) and ethene (C_2H_4).

Formation of Oxygen Molecule

Oxygen atom belongs to group VIA of the periodic table and it has 6 valence electrons in its outer shell. It needs two more electrons to achieve a stable octet electronic configuration. Each oxygen atom will share two of its outer electrons with another oxygen atom to form an oxygen molecule (O_2).Thus, two pair of electrons are shared between the two oxygen atoms to form a double covalent bond. The double covalent bond in an oxygen molecule is represented by a dot and cross diagram as shown in figure 4.7

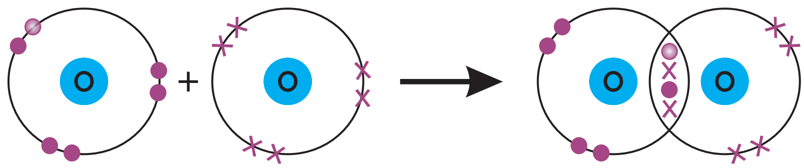
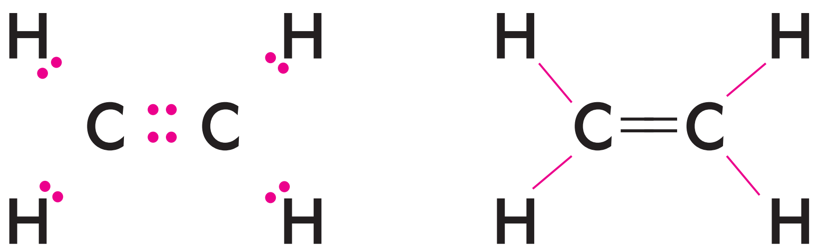


Fig 4.7 Formation of double covalent bond in oxygen molecule

The structural formula of an oxygen molecule is written

$$O=O$$

Another example of double covalent bond in ethene molecule can be represented as follows:



Triple Covalent Bond ($\equiv$)

A covalent bond which is formed by the mutual sharing of three bond pairs is called a triple covalent bond and it is represented by three short straight lines. For example, Molecule of nitrogen ($N\equiv N$) and ethyne ($CH\equiv CH$).

Formation of Nitrogen Molecule

Nitrogen is a non-metal. Each nitrogen atom has five electrons in its outer shells. Two nitrogen atoms will share three electrons to form three covalent bonds which is called triple covalent bond and form a nitrogen molecules (N_2). The triple bond in nitrogen molecule is represented by dot and cross diagram shown in figure 4.8

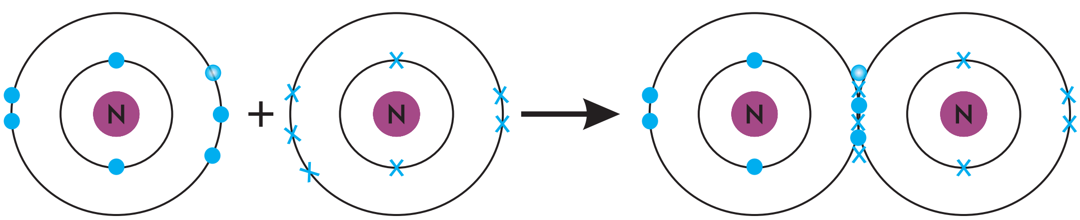


Fig 4.8 Formation of triple covalent bond in nitrogen molecule

The structural formula of a nitrogen molecule is:

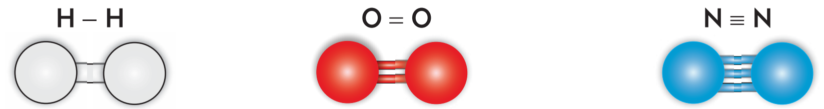
$$N \equiv N$$

Another example of triple covalent bond in ethyne molecule can be represented as follows:

$$H : C \vdots\vdots C : H \qquad H - C \equiv C - H$$

Thus, we can simply define three types of covalent bond as:

- Mutual sharing of two electrons between two atoms form a single covalent bond.
- Mutual sharing of four electrons between two atoms form a double covalent bond.
- Mutual sharing of six electrons between two atoms form a triple covalent bond.



Single bond   Double bond   Triple bond

4.3.3 Polar and Non-polar Covalent Bond

Covalent bonds are formed between two similar and dissimilar atoms.

For example, H - H, O = O, N≡N, H - Cl

Non-polar Covalent Bond

The covalent bond formed between identical atoms is called non-polar covalent bond. Both the identical atoms exert same force on the shared electron pairs. Non polar covalent bond in hydrogen molecule is shown below:

$$H\cdot \; + \; {}_{\times}H \longrightarrow H \overset{\bullet}{\underset{\times}{}} H$$

In the above example, each H atom has same electro negativity value of 2.1, therefore the covalent bond between them is considered as non-polar.

It means non-polar covalent bonds are formed when the electro negativities of the two atoms are equal.

Polar Covalent Bond

On the other hand, when different atoms share electron pair, both the atoms exert unequal forces on the shared electron pair. Such a covalent bond is called a polar covalent bond. For example, bond in HCl, H_2O, NH_3 are polar covalent bonds.

In the formation of polar covalent bond, one of the atoms will attract the shared electron pair more strongly than the other one. This atom will be called more electronegative atom. So, the more electronegative atom partially draws electron towards itself, this make it partially negatively charged δ^{-} and another atom partially positively charged δ^{+}.

For example, in Hydrogen chloride, Cl is more electronegative than the hydrogen. This causes the Cl atom to acquire a slight negative charge, and H atom aquire a slight positive charge due to electronegative difference. Thus, the bond between hydrogen and chlorine is called a polar covalent bond.

The compounds which have polar covalent bonds are called polar compounds.

Electronegative values determine whether a chemical bond will be ionic or covalent in nature. When the difference between electronegative values of two bonded atoms is more than 1.7, the bond will be purely ionic or electrovalent, whereas the difference is less than 1.7, the bond will be covalent. If the electronegative difference of bonded atoms is zero, the bond will be pure covalent or non-polar.

Do You Know?

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons.

Fluorine (the most electronegative element) is assigned a value of4.0, and values range down to cesium and francium which are the least electronegative at 0.7.

4.3.4 Coordinate Covalent Bond or Dative Covalent Bond

In the previous topic we have learned each atom contributes electron to form a covalent bond. However, covalent bond can be formed between two atoms even when only one of the atoms contributes both electrons constituting the covalent bond. Such a bond is called as a coordinate covalent bond or dative bond. Thus, we can define a coordinate covalent bond as:

The type of bond in which bond pair of electrons is contributed by one atom only, is called coordinate covalent or dative covalent bond.

Concept of donor and acceptor

The atom that donates the electron pair is called the donor and the other atom which accepts the electron pair is called acceptor. A coordinate covalent bond is represented by an arrow (→) pointing towards the atom which accepts the electron pair. A few examples of formation of a coordinate covalent bond are given as under:

Reaction between Ammonia and Hydrogen Chloride

The reaction between ammonia and hydrogen chloride involves the formation of a dative bond between N atom in NH_3 containing lone pairs and H^+ ion from HCl. When ammonia reacts with hydrogen ions (H^+) in an aqueous solution of an acid, the hydrogen ion is attracted to the lone pair and a coordinate covalent bond is formed.

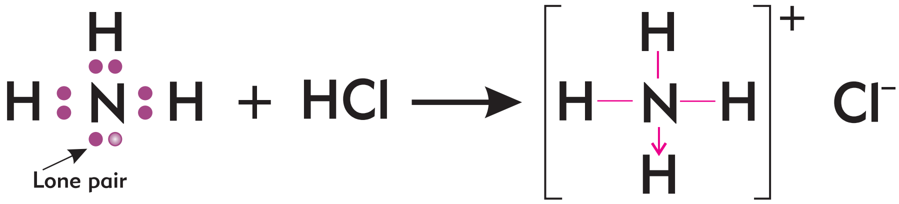
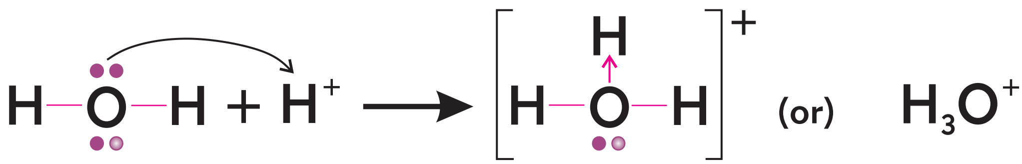


Fig 4.9 Reaction between Ammonia and Hydrogen Chloride

When hydrogen chloride dissolves in water, hydrogen ion is attracted to the lone pair of electrons which is available on oxygen and hydronium ion is formed as shownbelow:



Once a bond is formed, it is impossible to tell any difference between the dative covalent and ordinary covalent bonds. There is no difference between them in reality. The only difference between the two is a mode of formation. Due to their covalent nature of bond formation, the properties of these compounds are similar to those of covalent compounds.

Metallic Bond

Metallic bonds are formed by the attraction between metal ions and delocalized or "mobile" electrons within metal as shown in figure 4.10

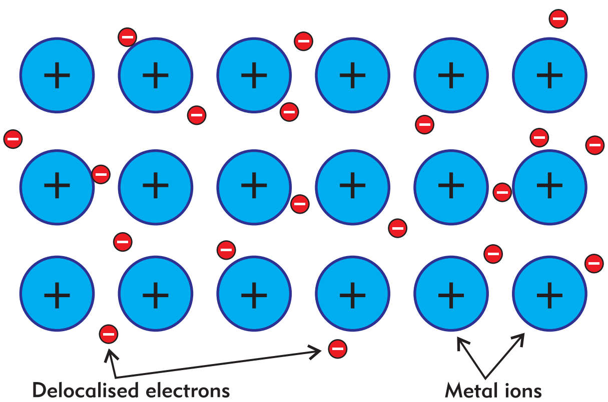


Fig 4.10 Diagramatic representation of metallic bonding

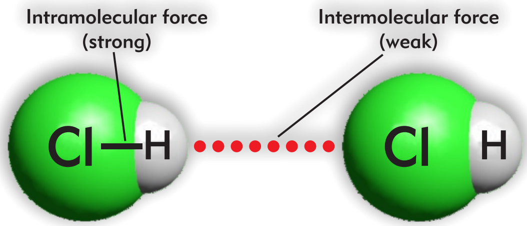
- Metal atoms loose the outer shell electrons and become positively charged ions and occupy a fixed position in a lattice.
- The outer shell electrons are free to move between the metal ions so are called delocalized electrons and move freely
- Thus the metal lattice structure shows positively charged ions surrounded by a delocalized outer electrons.

Test Yourself

- Magnesium is present in which group? How many electrons are in the outer shell of a magnesium atom?
- What is the charge of a magnesium ion and what is its symbol?
- Why Fluoride ion is not considered as Neon atom?
- Does anionic bond have a dipole?
- Describe the formation of anions for the following non-metals using dot cross structure.
  (a) Sulphur(atomic No.16) (b) Oxygen (atomic No.8)
- Where are valence electrons located, and why are they important?
- Why, noble gases do not react with other element to form compounds?
- Write the formation of cations for the following metals atoms using dot cross structure.
  (b) K (atomic No.19) (b) Al (atomic No.13)

4.4 INTERMOLECULAR FORCES

As we already discussed that some forces which hold the atoms together in a substance are called chemical bonds. Moreover, along with these strong bonding forces, weak forces are also created in between the molecules. These are called intermolecular forces. Thus, Intermolecular forces are defined as the set of all the forces that occur between two neighboring molecules. The bonding and intermolecular forces of hydrochloric acid are shown below:



These intermolecular forces are weaker than ionic and covalent bond. The interaction between intermolecular forces may be used to describe how molecules interact with each other. The strength or weakness of intermolecular forces determines the states of matter of a substance (e.g., solid, liquid, gas) and some of the physical properties (e.g., melting point, structure etc).

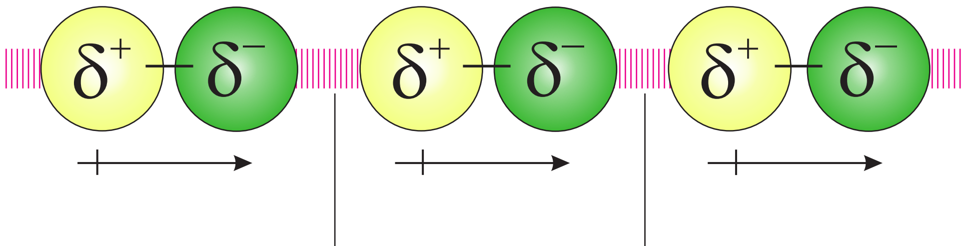
There are several types of intermolecular forces, but we will discuss two of them.

Do You Know?

Intra-molecular forces are forces between atoms within a single molecule. These forces are stronger than intermolecular forces as between H and Cl in HCl.

4.4.1 Dipole-Dipole Interaction

Dipole-Dipole interactions result when the two dipolar molecules interact with each other. When partially negative portion of one of the polar molecules is attracted to the partially positive portion of the second polar molecule, the electrostatic attraction is created between two molecules. These attractive forces are called Dipole-Dipole interactions and represented as below:



Dipole - Dipole Interactions

In the diagram, "δ" (read as "delta") means "slightly".

Example Dipole-Dipole Interaction

Dipole-dipole interaction can be seen in hydrogen chloride. Chlorine atoms are much more electro negative than hydrogen atoms. A partial negative charge is created on Chlorine and in turn a partial positive charge on hydrogen due to electronegativity difference.

$$\overset{\delta^+}{H}-\overset{\delta^-}{Cl}$$

When two molecules of hydrogen chloride come close to each other, the slightly negative end of one molecule is attracted to the slightly positive end of another molecule. These attractive forces are simply called dipole-dipole interaction as represented below:

$$\overset{\delta^+}{H}-\overset{\delta^-}{Cl}\text{-------}\overset{\delta^+}{H}-\overset{\delta^-}{Cl}$$

4.4.2 Hydrogen Bonding

Hydrogen bond is a type of dipole dipole interaction. When hydrogen forms polar covalent with more electronegative atom like Nitrogen (N), Oxygen (O), Fluorine (F), Chlorine (Cl), Sulphur (S), then hydrogen gets partially positive charge and other electronegative atom get partially negative charge. The interaction between partially positive charged hydrogen atom of one molecule with electronegative atom of other molecule is called Hydrogen bonding.

In molecules containing N-H, O-H or F-H bonds, the large difference in electronegativity between the H atom and the N, O or F atom leads to a highly polar covalent bond.

Because of the difference in electronegativity, the H atom bears a partial positive charge and the N, O or F atom bears a partial negative charge. (δ^+ and δ^- show slight charges).

$$\overset{\delta^-}{N}-\overset{\delta^+}{H} \qquad \overset{\delta^-}{O}-\overset{\delta^+}{H} \qquad \overset{\delta^-}{F}-\overset{\delta^+}{H}$$

The high partial positive charge on H atom enables to attract highly electronegative (N, O, or F) atom of the other molecule.

Example of Hydrogen Bonding

Consider the example of hydrogen fluoride. The fluorine atom is more electronegative. They tend to pull on the shared pair of electrons, creating a partial negative charge on itself and a partial positive charge on the hydrogen. The partial positive charge bearing hydrogen, then forms a bond with the electronegative atom of a neighboring molecule, while its electronegative element forms another bond with the positive hydrogen of another neighboring molecule. Therefore, several molecules combine by hydrogen bonding thus:

$$\text{- -}\ \overset{\delta^+}{H}-\overset{\delta^-}{F}\ \text{- -}\ \overset{\delta^+}{H}-\overset{\delta^-}{F}\ \text{- -}\ \overset{\delta^+}{H}-\overset{\delta^-}{F}\ \text{- -}$$

Hydrogen bonding

These intermolecular forces are extremely important in determining the properties of water, biological molecules, such as proteins, DNA etc. Synthetic material such as glue, paints and dyes are developed due to hydrogen bonding. Synthetic resins bind two surfaces together by hydrogen bonding or dipole-dipole interaction. Moreover, hydrogen bonding affects the physical properties of the molecules like melting and boiling point, density, solubility etc.

Test Yourself

- Why coordinate covalent bond is always a polar bond?
- Write down dot and cross structure of CCl_4
- Why does not a hydrogen atom form more than one covalent bond?
- Represent the formation of anions by the following nonmetals using electron 'dot' and 'cross' structure.
  (a) N (b) Br (c) P
- Why does a dipole occur in a molecule?

4.5 NATURE OF BONDING AND PROPERTIES

As discussed earlier the losing or gaining of electrons leads to ionic bonding, while the sharing of electrons leads to covalent bonding. The properties of the compounds depend upon the nature of bonds existing between them. Let us now discuss the effect of nature of bonding on the properties of compounds.

4.5.1 Ionic Compounds

Compounds having ionic bonds are called ionic compounds. The properties of ionic compounds relate to how strongly the positive and negative ions attract each other in an ionic bond. Most of the ionic compounds are in a solid or crystal form with strong electrostatic forces. Figure 4.11 shows the arrangement of Na^+ and Cl^- ions in NaCl. In crystal structure of sodium chloride each Na+ ion is surrounded by six Cl- ions. Similarly, each Cl- ions is surrounded by six Na+ ions.

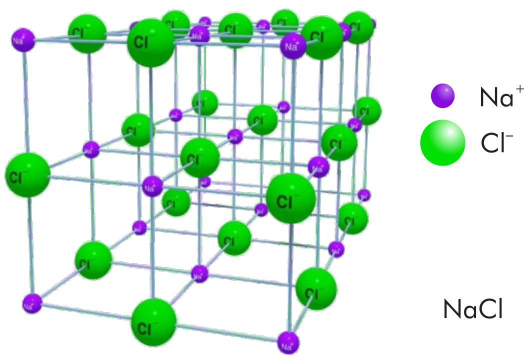


Figure 4.11 arrangement of ions in Solid crystal of NaCl

The ionic compounds exhibit the following properties:

i) Ionic compounds form crystals.
ii) Ionic compounds tend to be hard and brittle
iii) The large attracting forces result in a very stable structure. Therefore a lot of energy will be required to break these forces. So ionic compounds have high melting points. For example, melting point of NaCl is 801°C and boiling point 1413°C.
iv) Aqueous solutions of ionic compounds also conduct electricity. This is because when an ionic compound dissolves in water, the ions are free to move in aqueous solution.

v) Ionic compounds usually dissolve in polar solvent like water and are insoluble in non-polar solvents like oil, petrol, kerosene oil etc.

4.5.2 Covalent Compounds

As we know that covalent compounds are formed by mutual sharing of electrons between atoms. The bonding force of covalent bonds are generally weak as compared to ionic bond. The covalent compounds have following properties:

i) Covalent compounds usually exist as molecule i.e. H_2, O_2, H_2O but also can exist as crystals, examples include sugar crystals and diamond.
ii) The melting and boiling points of most covalent compounds are usually low.
iii) They are usually bad conductors of electricity.
iv) They are usually insoluble in water, but soluble in non-polar solvents like oil, petrol, kerosene, etc.

4.5.3 Polar and Non-polar Compounds

Polar and non polar both compounds differ in properties.

i) Nonpolar covalent compounds are generally insoluble in water while polar covalent compounds are soluble in water.
ii) Non-polar covalent compounds do not conduct electricity in the solid, molten or aqueous solution, but polar covalent compounds usually conduct electricity due to the formation of ions in water.
iii) Non-polar covalent compounds are soluble in non-polar solvent like petrol, benzene etc. While polar covalent compounds are insoluble in non-polar solvent.
iv) Few examples of polar covalent compounds are H_2SO_4, H_2O, HCl, HF, HBr, HI
v) Few examples of non-polar covalent compounds are CO_2, CH_4, C_2H_6.

4.5.4 Metal

Following are the different properties of metals such as:

- Metals are usually malleable and ductile.
- They are good conductor of electricity and heat due to the presence of delocalized electrons (mobile electrons).
- Melting and boiling points of metal are usually high as the atoms in metals are packed tightly.
- Metals have high densities.

Do You Know?

Malleable mean that metals can be hammered into different shapes and rolled into sheet.

Ductile is the property through which metal can be drawn into wires.

Test Yourself

- What do you understand about intermolecular forces between two molecules?
- Why metals are good conductor of electricity?

Society, Technology and Science

Uses of different synthetic adhesives like glue and epoxy resins

Synthetic adhesives like epoxy resins and glues are the substance that stick to the surface of the other objects. The material like plastic, wood, metal, ceramic glass and rubber etc at which glue is applied are called substrate. Epoxy adhesive is more expensive as compared to resin glue. Both are synthetic adhesives and require mixing before use, but epoxy hardens much faster than resin glue. We can use adhesive anytime to reattach the broken objects. For example, Poly vinyl acetate is a common white glue. It is used in book binding. Polyurethane glue is a flexible adhesive. It is used in fixing of soles to the bodies of shoes and wood working. Natural rubber bond to substrate on contact. It is used in self-adhesive envelopes. Conductive adhesive is commonly used in electronics to repair equipment. A mino resins are water soluble adhesive, they are used in bonding of layers in plywood. Epoxy glue contain epoxy. Its form strong bonds with glass, plastics, plywood, laminated boards and ceramic. Another use for epoxy resin is the decorative flooring applications. Commonly, epoxy resins are used where water resistance is required. Bridges, dams, power stations are also coated with epoxy resins.

Explain how Aircrafts, cars, trucks and boats parts are partially held together with epoxy adhesive (Analyzing)

The excellent adhesive properties of epoxy resins are due to the attractive forces between the epoxy resin and the surface of the substrate. One of the most common uses of epoxy resin is for adhesive purposes. For that purpose, epoxy resin is used in the construction of vehicles, trucks, boats and aircrafts. Its drying time is 6-30 minutes hardly.

Summary

- Every atom tries to achieve a noble gas configuration.
- Only outer most valence electrons are involved in bonding.
- Ionic bonding involves transfer of electrons.
- Metal reacts with non metal to form an ionic compound.
- Atoms which loose electron(s) form positive ions. Atoms which gain electron(s) form negative ion.
- In an ion, the number of electrons is different from the number of protons.
- Ionic bonding is commonly formed between elements of group IA or II A and groups VI A or VII A.
- Covalent bonding involves sharing of electrons and form molecules.
- The sharing of three pairs of electrons between two atoms is called a triple bond.
- Metal tends to lose valence electrons to form positively charged ions (cations).
- Non metals usually gain electrons to form negatively charged ions (anions)
- Common covalent molecules are water H_2O, Methane CH_4, Ammonia $NH_3$and carbon dioxide CO_2.
- A co-ordinate bond also called a dative covalent bond.
- A covalent bond can be polar or non-polar. But coordinate bond is only polar in which both electrons come from the same atom.
- The sharing of two pairs of electrons between two atoms is called a double covalent bond.
- A hydrogen bond is a partially electrostatic attraction between a hydrogen (H) which is bound to a more electronegative atom such as nitrogen (N), oxygen (O), or fluorine (F), and another adjacent atom bearing a lone pair of electrons
- When slightly negative end of polar molecule is weakly attracted to the slightly positive end of another molecule then such attracting forces are called dipole-dipole interactions.

EXERCISE

SECTION- A: MULTIPLE CHOICE QUESTIONS

Tick Mark (✓) the correct answer

1. An example of ionic compound is:
   (a)H_2 (b)CH_4 (c) N_2 (d)NaCl
2. Interaction between highly electron deficient hydrogen and highly electronegative atom is called
   (a) covalent bond (b) ionic bond
   (c) hydrogen bond (d) metallic bond
3. Two fluorine atoms share one electron each in their outermost shell to achieve electronic configuration of:
   (a) Xe (b) Ar
   (c) Kr (d) Ne
4. Number of electrons lost by atoms of group IIIA equals:
   (a) 1 (b) 2
   (c) 3 (d) 4
5. Atom which loses two electrons from its outer shell to form ion is:
   (a) oxygen (b) potassium
   (c) magnesium (d) carbon
6. In NaCl crystal lattice each Na^+ ion is surrounded by:
   (a) 6 Cl^- ions (b) 6 Na^+ ions
   (c) 8 Cl^- ions (d) 12 Cl^- ions
7. At room temperature most of ionic compounds are:
   (a) amorphous solids (b) crystalline solids
   (c) liquids (d) gases
8. Tendency of atoms to acquire eight electrons in their valence shell is:
   (a) octet rule (b) duplet rule
   (c) triplet rule (d) none of above
9. When one atom forms cation by losing electron and other forms anion by accepting that electron then bond form between them is:
   (a) Covalent bond (b) Ionic bond
   (c) coordinate covalent bond (d) hydrogen bond

10. Nobel gases are stable because they contain:
    (a) 4 electrons in valence shell (b) 6 electrons in valence shell
    (c) 8 electrons in valence shell (d) 10 electrons in valence shell
11. Bond which involve 3 shared electron pairs is a :
    (a) double covalent bond (b) single covalent bond
    (b) triple covalent bond (d) none of above
12. A non-metal atom form anion by
    (a) loses of electrons (b) gain of electrons
    (c) loses of protons (d) gain of protons
13. When two identical atoms share electron pairs and exert same force on each other than
    bond form is:
    (a) non-polar covalent bond (b) polar covalent bond
    (c) double covalent bond (d) coordinate covalent bond
14. Synthetic resins are used on places where:
    (a) electric resistance is required (b) water resistance is required
    (c) adhesion is required (d) friction is required
15. Oxygen belongs to group VIA so number of electrons in its valence shell are:
    (a) 4 (b) 2
    (c) 6 (d) 8
16. Electron pairs which are not shared by atoms are called:
    (a) electron pairs (b) lone pairs
    (c) bond pairs (d) shared pairs
17. Strength of intermolecular forces from ionic or covalent bond is:
    (a) Weaker (b) stronger
    (c) equal (d) none of above
18. Ionic crystals have:
    (a) high melting points (b) moderate melting points
    (c) low melting points (d) none of above
19. Bond formed by mutual sharing of electron is:
    (a) ionic bond (b) coordinate covalent bond
    (c) covalent bond (d) metallic bond

20. Which of the following diagram shows atoms are bonded with same electro negativity?

(a) A —:— B   (b) A ——: B

(c) A :—— B   (d) A —:—— B

SECTION- B: SHORT QUESTIONS:

1. Draw dot and cross diagrams to show how different types of chemical bonds are formed when fluorine reacts with
   (a) Hydrogen   (b) potassium
2. What is meant by octet and duplet rule?
3. Can you draw an ion which is formed by the atom losing three electrons?
4. How oxygen forms an anion?
5. What is the difference between lone pair and bond pair?
6. Explain why table salt has a very high melting point.
7. How is electronegative value determined the formation of chemical bond?
8. Why it is easy for magnesium atom to lose two electrons?
9. Atoms of metallic elements can form ionic bond, but they are not very good to form covalent bonds. Why?
10. How does an ion differ from an atom?
11. Describe dipole-dipole forces.
12. Write uses of adhesive material.
13. Why Intermolecular forces are weaker than intera molecular forces?
14. Write characteristics of metallic bond.
15. Ionic bonds are strong and hard to break but why most of the covalent compounds have low melting and boiling points.
16. Write down the characteristics of ionic compounds.
17. Why ionic compounds are solid?
18. How is hydrogen bonding affecting the physical properties of compounds?
19. Complete the chart:

Atomic Number	Number of protons	Number of electrons	Electronic configuration	Number of valence electrons
11	11	11	2, 8, 1	1
12				
13				
14				
15				
16				

SECTION- C: DETAILED QUESTIONS:

1. Define ionic bond. Discuss the formation of sodium chloride (NaCl).
2. Explain element attain stability?
3. Describe the formation of a covalent bond between two nonmetallic atoms. Explain single, double and triple covalent bond with examples.
4. How are electrons arranged in molecular compound? Draw electron dot and cross structures for the following atoms.
   (a) H_2O (b) N_2 (c) CH_4 (d) C_2H_2 (e) Cl_2 (f) H_2
5. Define metallic bond. How are metallic bonds are formed?
6. What is coordinate covalent bond? Explain with two examples.
7. What do you understand about ionic character of covalent bond?
8. Differentiate the properties of polar and non-polar covalent compounds.
9. Explain the importance of glues and epoxy resins in our society.

Chapter 5

PHYSICAL STATES OF MATTER

Time Allocation

Teaching periods	= 12
Assessment period	= 3
Weightage	= 12

Major Concepts

5.1 Gaseous State
5.2 Laws Related to Gases
5.3 Liquid State
5.4 Solid State
5.5 Types of Solids
5.6 Allotropy
5.7 Plasma State
5.8 Bose Einstein Condensate

STUDENTS LEARNING OUT COMES (SLO'S)

Students will be able to:

- Understand effect on the pressure of gas by change in the volume and temperature.
- Compare the physical states of matter with reference of intermolecular forces present between them.
- Account for pressure volume changes in a gas using Boyle's Law.
- Account for temperature volume changes in a gas using Charles's Law.
- Summarize the properties of liquids like evaporation, vapor pressure and boiling point.
- Explain the effect of temperature and external pressure on vapor pressure and boiling point.
- Describe physical properties of solids (boiling and melting points).
- Differentiate between amorphous and crystalline solids.
- Explain the allotropic forms of solids.
- Define the plasma with help of examples.
- Define Bose Einstein Condensate with help of example.

Introduction:

As we know that matter is physical material of universe. It is defined as anything which has mass and occupy space. States of matter differ in some observable properties. A gas has no fixed volume or shape it can easily be compressed and expanded. A liquid state has no fixed shape but has fixed volume it cannot be compressed easily. A solid has a definite shape and volume, it cannot be compressed. In addition to the above three states, there are two more states of matter named as plasma state and Bose Einstein condensate. The different physical states of matter are due to arrangement of molecules and intermolecular forces.

Gaseous State:

The gaseous state molecules are lying away from one another, this assumption was proposed by Boltzmann, Maxwell, Kelvin. They explained the behavior of gases according to their kinetic molecular theory. Gaseous state shows following characteristics

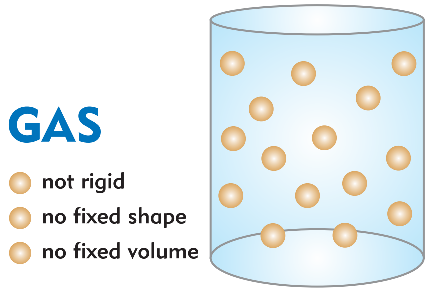


Fig 5.1

- The molecules in gases are widely separated from each other.
- They have negligible volume.
- The gas molecules are in constant random motion.
- The gas molecules moves in straight line until they collide with each other or wall of container.
- On collision molecules do not lose energy because their collision is perfectly elastic.
- Pressure Produced when molecules collide with the wall of container.
- There are no attractive and repulsive forces between molecules.

5.1 PROPERTIES OF GASES

The kinetic molecular theory explains the behavior of gases such as Diffusion, Effusion, Pressure, compressibility, Mobility and Density which are defined below:

5.1.1 Diffusion:

The Diffusion is defined as spontaneous mixing of molecules by random motion and collision to from a homogenous mixture. Gases are rapidly diffusible and rate of diffusion depend upon the molecular mass of the gases. Lighter gases diffuse rapidly than heavier gases as H_2 diffuses four times more than O_2.

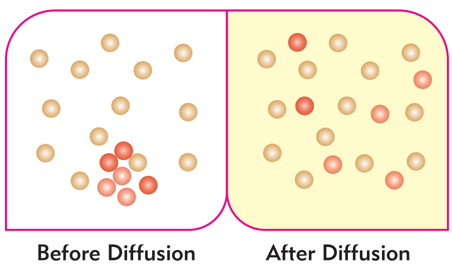


Fig 5.2 Diffusion

Diffusion is movement of particles from an area of higher concentration to lower concentration. The rate of this movement depends upon temperature, viscosity of the medium and the size or mass of the particles. Diffusion results in the gradual mixing of materials, and eventually it forms a homogeneous mixture.

For Examples:

- You can smell perfume because it diffuses into the air and makes its way into your nose.
- Smoke diffuses into the air.
- Flower smell, garbage stink or body odor.

5.1.2 Effusion:

The Effusion is escaping of gas molecules through a tiny hole into a space with lesser pressure. Effusion depends upon molecular masses of gases .Lighter gases effuses rapidly than heavier gases. For process of effusion the diameter of hole must be smaller than the molecule's mean free path.

For Examples:

- Leakage of air through tyre pin hole.
- Leakage of helium through gas balloons.

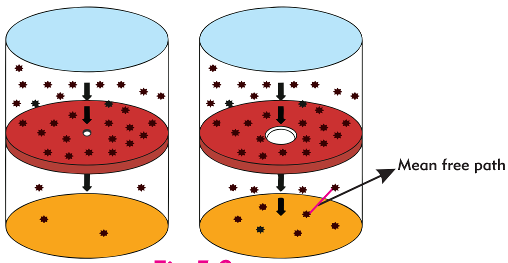


Fig 5.3

Do you know?

Mean free path is the average distance that a gas particle travels between successive collisions with other gas particles.

5.1.3 Pressure

The force exerted by gaseous particles per unit area is called gas pressure. It can be expressed mathematically as.

Pressure = Force / Area or $P = F / A = N/m^2$

The S.I unit of force is Newton (N) and unit of area is m^2, hence pressure has S.I unit of Nm^{-2}. It is also known as Pascal (Pa). 1 Pascal = $1 Nm^{-2}$.

The molecules of gases are in continuous motion,

Do you know?

Pressure exerted by the atmosphere at the sea level is called atmospheric pressure. It is defined as the pressure exerted by a mercury column of 760 mm height at sea level.

1atm = 760mm of Hg = 760 torr

1atm = 101325 pascal

But the pressure is developed by the collisions of molecules of gas with the walls of container. Barometer is used to measure the atmospheric pressure and manometer is used to measure the pressure of gases in the laboratory.

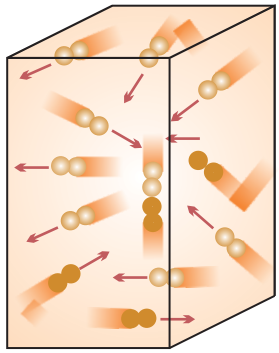


Fig 5.4

5.1.4 Compressibility

The capacity of something to be flattened or reduced in size by pressure is called compressibility. The gases are highly compressible due to larger spaces between their molecules. When gases are compressed, the molecules come closer to one another and occupy less volume as compared to the volume of uncompressed state.

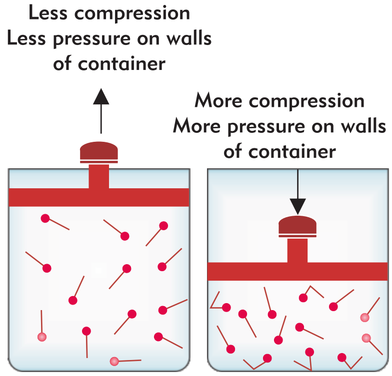


Fig 5.5

5.1.5 Mobility:

The ability to move freely is known as mobility. As the gas molecules are in continuous motion they can move fast due to high kinetic energy. The molecules move freely in the empty space .This mobility is responsible to produce a homogenous mixture of gases.

5.1.6 Density:

The Density is degree of compactness or closeness of a molecules. Gases have low density because of light mass and more volume occupied by the gas molecules. Gas density is expressed in grams per dm^3. Gases are less denser than liquids. The density of gases can be increased by cooling.

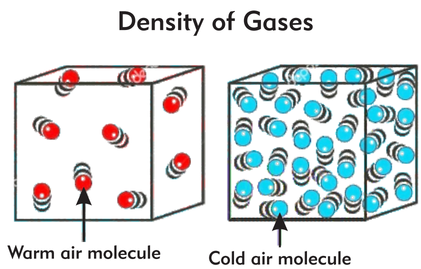


Fig 5.6

Table 5.1 Density of different gases

Gas	Chemical formula	Density Kg/m^3
Oxygen	O_2	1.407
Chlorine	Cl_2	3.120
Carbon dioxide	CO_2	1.935
Hydrogen	H_2	0.088
Nitrogen	N_2	1.232
Helium	He	0.176

Test Yourself

- Why gases diffuse rapidly, Explain?
- Why density of gases increases on cooling?
- Explain that Effusion depends on mean free path.

5.2 LAWS RELATED TO GASES

The properties of gases are governed by the following laws.

5.2.1 Boyle's Law:

In 1662 Robert Boyle proposed gaseous law about the relationship between volume and pressure of gas at constant temperature, Boyles law states that "The volume of a given mass of a gas is inversely proportional to applied pressure, at constant temperature"

Mathematical Representation of Boyle's Law:

According to Boyle's law the volume (V) of a given mass of a gas decreases with the increase of pressure (P) at constant temperature.

$$V \propto 1/P \quad \text{Or} \quad V = K/P \quad \text{where } K = \text{is the constant}$$

$$PV = K$$

The product of pressure and volume of a fixed mass of a gas is constant at a constant temperature.

If $P_1V_1 = K$ Then $P_2V_2 = K$

Where P_1 = initial pressure P_2 = final pressure

V_1 = initial volume V_2 = final volume

As the both equations have same constant therefore their variable are also equal to each other so $P_1V_1 = P_2V_2$

This equation establish a relationship between pressure and volume. The relationship between volume and pressure can be well defined from following figure (5.7). Where given mass of a gas at constant temperature shows increase in volume by decrease in pressure. On the other hand increase in pressure decreases volume. But the product of pressure and volume is constant in both cases. (Table 5.2)

Table 5.2 Boyl's Relationship between pressure and volume

P (Change in pressure)		V (Change in Volume)		K (Constant pressure)
1.0	x	4	=	4
2.0	x	2	=	4

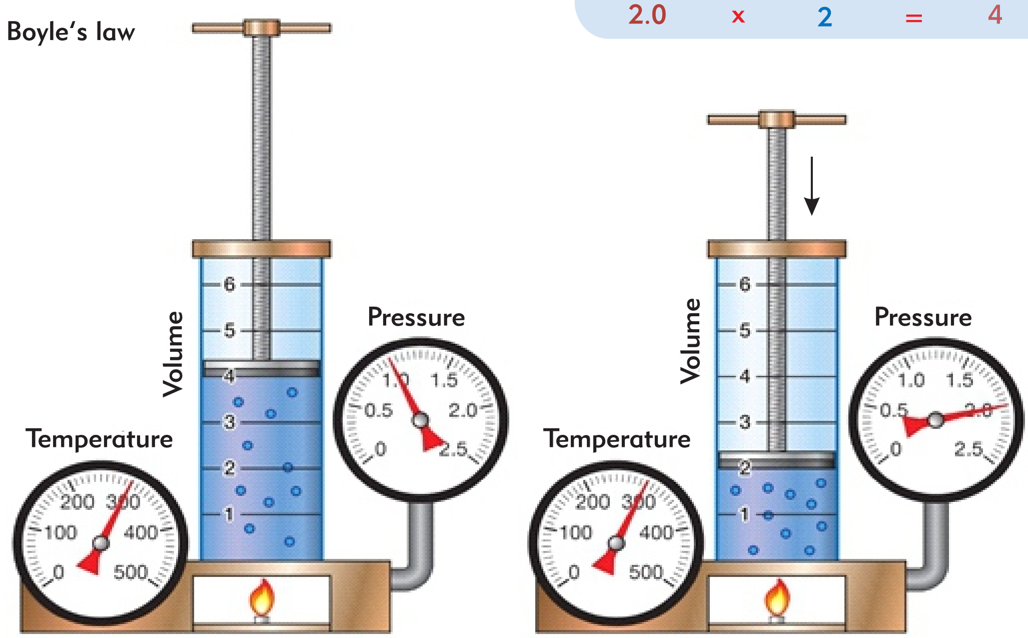


Fig 5.7

Examples 5.1

The pressure of a sample gas is 3 atm and the volume is 5 liters. If the pressure is reduced to 2 atm, what will be the new volume?

Data:

V_1 = 5 liter
P_1 = 3 atm
P_2 = 2 atm
V_2 = ?

Solution:

$$P_1V_1 = P_2V_2$$

$$V_2 = \frac{P_1V_1}{P_2}$$

$$V_2 = \frac{3\times5}{2}$$

$$V_2 = \frac{15}{2}$$

$$V_2 = 7.5 \text{ Litre}$$

The volume will be 7.5 liters.
The volume is increased by decreasing the pressure.

Examples 5.2

The 700 cm^3 of a gas is enclosed in a container under a pressure of 650 mm of Hg. If the volume is reduced to 350 cm^3 , what will be the pressure then?

Data:

V_1 = 700cm^3
P_1 = 650mm Hg
V_2 = 350cm^3
P_2 = ?

Solution :

$$P_1V_1 = P_2V_2$$

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

By putting the values $P_2 = 650 \times 700/350 = 1300$mm of Hg

Thus pressure is increased by decreasing volume.

5.2.2 Charles' Law

In 1787 French scientist J.Charles proposed his law to explain the relationship between volume and temperature keeping the pressure constant. He states that "the volume of a given mass of a gas is directly proportional to the absolute temperature if the pressure is kept constant".

Mathematical Representation of Charles' Law:

According to Charles law if temperature of a gas is increased, its volume will also increase.

Mathematically it is represented as:

$$V \propto T$$

Or

$$V = KT$$

$$\frac{V}{T} = K$$

Where K is proportionality constant. When temperature increases the volume also increases.

For example, if we double the temperature from 300 K to 600 K, at constant pressure, the volume of a fixed mass of the gas will become double (Fig 5.8).

Do you know?

Absolute Temperature Scale

Lord Kelvin introduced absolute temperature scale or Kelvin scale.it starts from zero (0K which is equal to -273°C). It is the temperature at which an ideal gas would have zero volume and known as absolute zero. As Celsius and kelvin scales have equal degree range therefore zero kelvin is equal to -273 Celsius, and 273 kelvin is equal to zero Celsius.

Conversion of kelvin temperature and Celsius temperature are vice versa as follows:

$(T)K = (T)°C + 273$

$(T)°C = (T)K - 273$

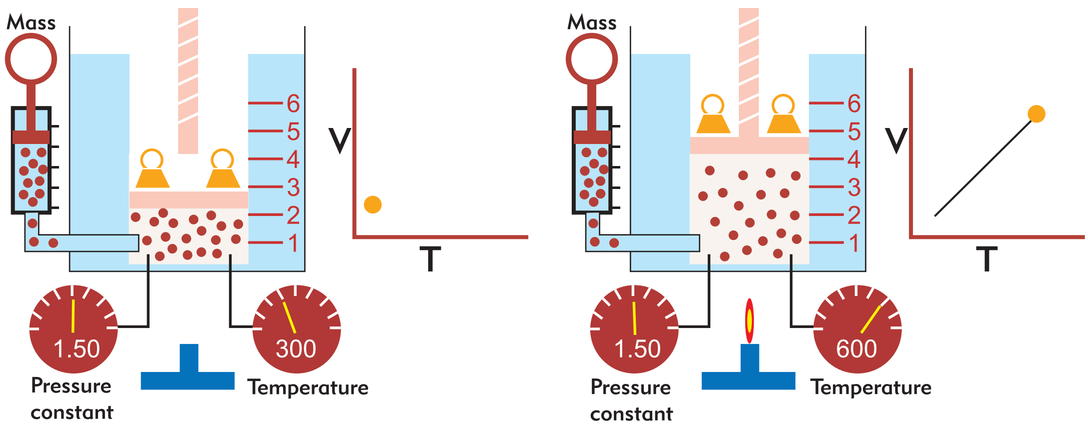


Fig 5.8

Imagine a gas at a certain temperature (T_1) and has volume (V_1). If change in the temperature (T_1) to a new value (T_2), occur then volume (V_1) changes to a new value (V_2). We can use Charles's law to describe both sets of conditions:

Initial State : $$\frac{V_1}{T_1} = K$$

Final State : $$\frac{V_2}{T_2} = K$$

The constant, k, is the same in both cases, therefore

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

Do you know?

Always convert temperature scale from $^{\circ}C$ to K while solving a problem.

$K = 273 + {}^{\circ}C$

Examples 5.3

A 600 ml sample of gas is heated from 27 °C to 77 °C at constant pressure. What is the final volume?

Data:

$T_1 = 27^{\circ}C = 27 + 273\ K = 300\ K$

$T_2 = 77^{\circ}C = 77 + 273\ K = 350\ K$

$V_1 = 600ml$

Solution:

By using the equation

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

or

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

By putting the values in equation

$$V_2 = \frac{600 \times 350}{300}$$

$$V_2 = 700\ ml$$

The Volume will become 700 ml, which shows increase in volume with raising the temperature.

Examples 5.4

A sample of Hydrogen gas has a volume of 350 cm^3 at 40°C. If gas is allowed to expand up to 700 cm^3 at constant pressure. Find out its final temperature?

Data:

$T_1 = 40\ ^\circ C = 40 + 273\ K = 313\ K$

$V_1 = 350\ cm^3$

$V_2 = 700\ cm^3$

$T_2 = ?$

Solution:

By using the Charles' law equation

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

or

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

By putting the values

$$T_2 = \frac{700 \times 313}{350}$$

$$T_2 = 626\ K$$

The expansion in volume is due to increase in temperature.

Test Yourself

- Which variables are kept constant in Boyle's law?
- When a gas is allowed to expand, what will be the effect on its temperature?
- What is absolute zero temperature?
- Can you reduce temperature of a gas by increasing its volume?

5.3 LIQUID STATE

The liquid state is the intermediate between gaseous and solid states. According to their kinetic molecular theory liquid state shows following chractarstics.

- The molecules of a liquid are randomly arranged like gases.
- The molecules of liquids have less kinetic energy than gases.

- The molecules of liquids are fairly free to move.
- The Liquids has no definite shape but assumes the shape of container.
- The Boiling point of liquids depends on the external atmospheric pressure.
- The Liquids are denser and not compressible like gasses.

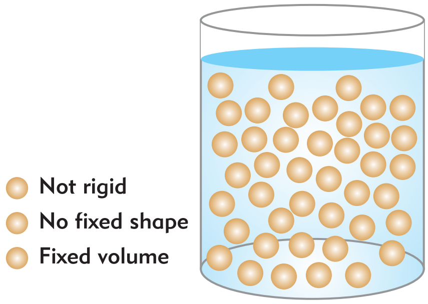
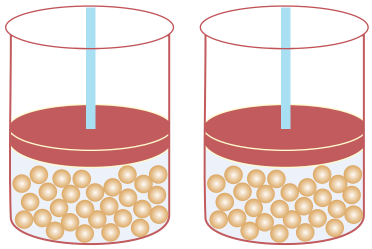


Fig 5.9



Liquid particles are very close together, so they are not compressed easily.
Fig 5.10

Properties of Liquids:

5.3.1 Evaporation

The process by which a liquid changes to a gas phase is called evaporation. Evaporation is endothermic reaction in which heat is absorbed.

For example clothes dry under the sun due to evaporation, in this process water is converted from liquid state into vapours by acquiring Heat.

Fig 5.11

$$\text{Water (liquid)} \longrightarrow \text{vapour (gas)}$$

The molecules of liquids are in continuous motion they collide with each other but all the molecules do not have same kinetic energy. Majority of the molecules have average kinetic energy and few have more than average kinetic energy. The molecules having more average kinetic energy overcome the attractive forces among the molecules and escape from the surface by evaporation. It is directly proportional to temperature and increase with the increase in temperature.

The evaporation is also considered as cooling process because when high kinetic energy molecules escape in the air in the form of vapours by taking energy from lower molecules as a result the energy of remaining molecules falls down. To compensate this deficiency of energy molecules absorb energy from the surrounding, due to which temperature of surrounding decreases and cooling occurs.

Factors Affecting Evaporation:

I. Surface Area: The evaporation is a surface based process. Greater the surface area greater is evaporation. For example water left in bowl evaporate slowly than water left in a large tub. A saucer is used to cool the tea quickly than tea cup.

ii. Temperature: The rate of evaporation increases with the increase in temperature. Because temperature increases the kinetic molecular energy which overcome the intermolecular forces and makes evaporation rapidly. For example clothes dry quickly in a sunny day than in a cloudy day.

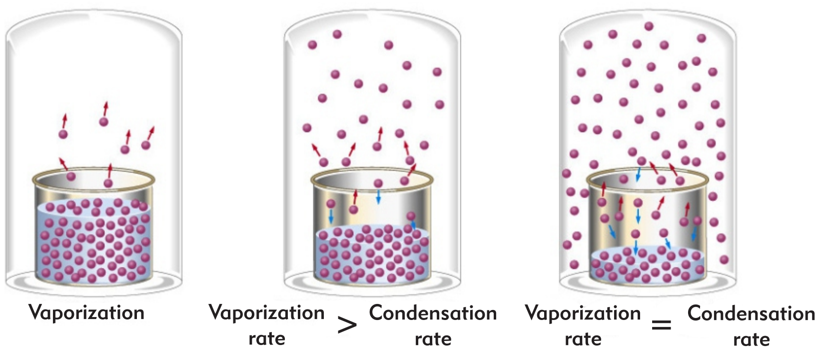
iii. Intermolecular Forces: The rate of evaporation increases with weak intermolecular forces .If intermolecular forces are stronger evaporation is lesser. For example perfume have weak intermolecular forces than water therefore it evaporates quickly.

5.3.2 Vapour pressure

The pressure exerted by vapours at equilibrium on its pure liquid state at a particular temperature is called Vapour Pressure. The equilibrium is a state when rate of evaporation is equal to rate of condensation but in opposite directions.

$$\text{Liquid} \underset{\text{Vaporize}}{\overset{\text{Condensation}}{\rightleftharpoons}} \text{Vapours}$$

Vapour pressure takes place in a closed system on heating because in open system or open surface liquid molecule evaporates and mix-up with air.



When a liquid is heated in a closed container evaporated molecules start gathering over the liquid surface. Initially the vapours condense slowly to liquid .After sometimes condensation process increases and a stage reaches when the rate of evaporation become equal to rate of condensation. At that stage, the number of molecules evaporate will be equal to the number of molecules condensate (coming back) to liquid . At this point, pressure exerted by the vapours on liquid is called vapour pressure. The unit of pressure is expressed in mm of Hg, atmosphere, torr or newton.

Factors Affecting Vapour Pressure:

i. **Nature of Liquid:** The vapour pressure depends upon the nature of liquids .Polar liquids have low vapour pressure than nonpolar liquids at the same temperature. It is because of strong intermolecular forces of molecules and high boiling point of polar liquids. For example water (polar liquid) has less vapour pressure than petrol (non polar liquid).

ii. **Size of Molecules:** The vapour pressure is more in small size molecules because small sized molecules evaporate easily and exert more vapour pressure. For example hexane (C_6H_{14}) has small sized molecules as compared to decane ($C_{10}H_{22}$), due to this hexane evaporate rapidly and exert more pressure.

iii. **Temperature:** The vapour pressure increases with raise in temperature. The average kinetic energy of molecules increases with temperature which causes increase in vapour pressure. For example vapour pressure of water at 0° is 4.58 mm Hg while at 100°C it increases up to 760mm Hg.

5.3.3 Boiling Point

The temperature at which vapour pressure of a liquid become equal to atmospheric pressure is called boiling point of that liquid. When the liquid is heated, bubbles begin to form throughout its volume. The bubbles (contain vapour pressure) which being lighter than the liquid, rise to surface and burst. The vapour pressure in bubble is equal to atmospheric pressure. All bubbles containing vapours will rise to the surface of the liquid and burst into air. It appears that water is boiling. The boiling point varies with the atmospheric pressure.

Factors Affecting Boiling Point:

i. **Atmospheric Pressure:** The boiling point is directly proportional to atmospheric pressure. Boiling point can be increased by increasing atmospheric pressure. For example working of pressure cooker.

ii. **Nature of Liquid:** The boiling point depends upon the nature of liquid as polar liquids have high boiling point than nonpolar liquids, because polar liquids have stronger intermolecular forces than nonpolar liquids. Boiling points of few liquids are given in table 5.3.

iii. **Intermolecular forces:** The intermolecular forces play very important role in the boiling points of liquids. Substances having stronger intermolecular forces have high

Table 5.3 Boiling Points of Some Common Liquids

S.No	Liquids	Boiling (°C)
1	Diethyl ether	34.6
2	Ethyl alcohol	78
3	Water	100
4	n-octane	126
5	Acetic acid	118
6	Mercury	356
7	Sulphuric Acid	330
8	Bromine	58.8

boiling points, because such liquids attain a level of vapour pressure equal to external pressure at high temperature.

5.3.4 Freezing Point:

The temperature at which the vapour pressure of a liquid state becomes equal to the vapour pressure of its solid state is known as Freezing Point of a liquid. At this temperature liquid and solid coexist in dynamic equilibrium.

Factors Affecting Freezing Point:

The freezing point depends upon the temperature and intermolecular forces. In comparison molecules with stronger intermolecular forces are pulled together to form a solid at higher temperature. Due to this they show high freezing point. Molecules with lower intermolecular forces solidify on more lower temperature.

Freezing points of few liquids are given in table(5.4)

Table 5.4 Freezing Points of Some Common Liquids

S.No	Liquids	Freezing point℃
1	Benene	5.12
2	Ethyle alcohol	-114
3	Water	0.0
4	Acetic acid	16.6
5	Mercury	-38.83
6	Sulphuric Acid	10.6
7	Bromine	-7.2

5.3.5 Diffusion:

The diffusion is defined as spreading out of the molecules throughout the vessel. The liquids diffuse less rapidly than gases.

As the molecule of liquid are in cluster and bounded with intermolecular binding forces, the liquid molecules roll over one another and are in continuous motion. They move from higher concentration to lower concentration and mix up with molecules of others liquids and may form homogenous mixture.

For example when few drops of ink are dropped in water filled flask, the molecules of ink move around and after a while spread in whole of flask, thus diffusion takes place as shown in figure 5.12

Fig 5.12

Factors Affecting Diffusion:

i. Inter molecular forces : liquids have weaker intermolecular forces than solid due to this they diffuses faster than solid but less rapidly than gases.

ii. Size of Molecules: Diffusion depends upon size of molecules small size molecules diffuses rapidly than bigger one. For example diffusion of methanol of (CH_3OH) is higher than ethanol (C_2H_5OH).

iii. Shape of molecules: Molecules with irregular shape diffuses slowly while regular shaped molecules diffuses faster because they can easily slip over and move faster.

iv. Temperature: Diffusion increases by increasing temperature because at high temperature intermolecular forces becomes weak due to high kinetic energy of the molecules.

5.3.6 Mobility:

The mobility is ability to move freely. The molecules in a liquid move freely, due to free movement they can adjust their shape in a container. Due to this reason liquids can flow.

Factors Affecting Mobility:

I. Temperature: Mobility increases by increasing the temperature. When temperature increases in a liquid movement of molecules increases accordingly.

ii. Inter molecular forces: Mobility increases by decrease in intermolecular forces. The liquids which have strong intermolecular forces show less mobility.

5.3.7 Density:

The Density of liquid is defined as mass per unit volume. Liquids are denser than gases due to closeness of molecules and small intermolecular spaces. As the molecules of liquids are close to one another, they have strong intermolecular forces they cannot expand freely and shows definite volume which makes liquids denser than gases.

Mathematically it can be expressed as $D = \frac{M}{V}$ $\begin{pmatrix} M = \text{mass} \\ V = \text{volume} \end{pmatrix}$

Factors Affecting Density:

I. Temperature: Density of liquids is slightly affected by the temperature as by increasing temperature liquids increase their volume which decreases density. Different densities of water at different temperature is given in Table 5.5

Table 5.5 Densities of water at different temperature

T (°C)	Density (g/cm^3)
0	0.99984
30	0.99565
60	0.98320
90	0.96535

ii. Pressure : Density of liquids is slightly affected by pressure .Increase in pressure on liquids increases the density, but liquids are not readily compressed due to this density change is negligible.

Test Yourself

- Why increase in temperature causes increase in the process of evaporation?
- Why rate of diffusion in liquids is less than gases? Prove with the help of examples.
- How equilibrium state involved in vapour pressure of liquids in close system? Explain how evaporation causes a cooling effect?
- How boiling point of a substance is affected by atmospheric pressure?

5.4 SOLID STATE

Solids have fixed shape and volume. The molecules in solids are extremely closer to each other, without having space between them. According to kinetic molecular theory solid state shows following characteristics.

- The molecules in solids are closely packed due to stronger forces of attraction.
- The molecules are unable to move freely as they have little space between them.
- The molecules can vibrates and rotate in their fixed position.
- Solids have definite shape and definite volume.
- Pure solids have sharp melting point.

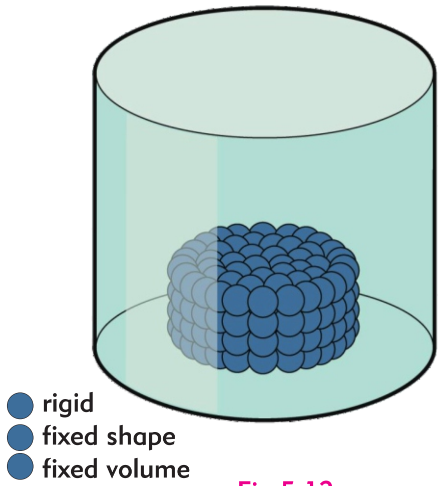


Fig 5.13

Properties of Solid

5.4.1 Melting Point:

The temperature at which a solid starts melting and coexist in equilibrium with liquid state is called melting point. when temperature is raised the kinetic energy of molecules increases. When continuous heat is supplied to solid, molecules leave their fixed position and become mobile, at this stage solid starts melting.

The melting points of different solids are given in table 5.6.

Table 5.6 Melting points of solids

S.No	Solids	Melting point °C
1	Sugar	185
2	Chocolate	36
3	Mercury	-38.83
4	Sodium chloride	801
5	Water	0

5.4.2 Rigidity:

The molecular arrangement of solids is closely packed due to this solids are not mobile. They exhibit vibration at fixed positions. Therefore solids are rigid in their structure.

5.4.3 Density:

The solids are typically denser than a liquid or a gas. Molecules in solid are more tightly packed together because of the greater intermolecular forces. Due to this reason solids have highest densities among three states of matter. Some densities are shown in table 5.7.

Table 5.7 Densities of solids

S.No	Solids	Density (g/cm^3)
1	Aluminum	2.70
2	Iron	7.86
3	Gold	19.3
4	Sodium chloride	2.16
5	Sugar	1.59

5.5 TYPES OF SOLIDS

The solids are classified on the basis of arrangement of molecules. There are two types of solids which are as follows.

i. Crystalline solids
ii. Amorphous solids

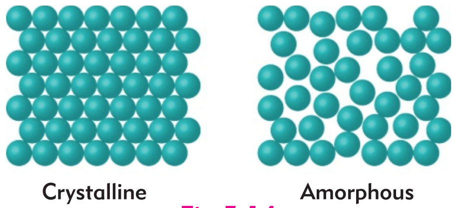


Fig 5.14

5.5.1 Crystalline Solids:

The solids in which molecules are arranged in definite three dimensional geomatrical pattern are called crystalline solids. The arranged molecules in solids are bounded biplane surface called faces which intersect each other at a particular angle. The melting point of crystals are sharp. Sodium chloride and diamond are common examples of crystalline solids.

Fig 5.15 sodium chloride

Fig 5.16 Diamond

5.5.2 Amorphous solids:

The solids in which molecules are not arranged in geometrical pattern are called amorphous solids. They do not have sharp melting point. Plastic ,rubber and glass are examples of amorphous solids.

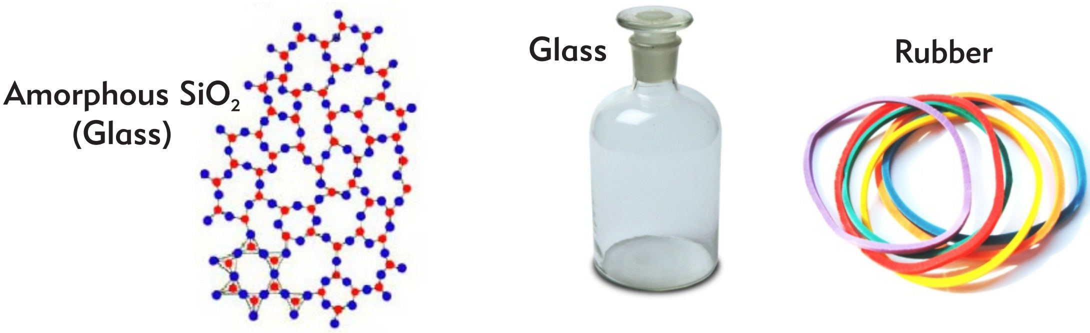


Fig 5.17

5.5.3 Difference between Amorphous and Crystalline Solids

Amorphous Solids	Crystalline Solids
They don't have definite geometrical shape.	They have characteristic geometrical shape.
Amorphous solids do not have particular melting point. They melt over a wide range of temperatures.	They have sharp melting point.
Amorphous solids are isotropic.	Crystalline solids are anisotropic.
Amorphous solids are unsymmetrical.	Crystalline solids are symmetrical.
Amorphous solids do not break at fixed cleavage planes.	Crystalline solids break along particular direction at fixed cleavage planes into small pieces of same shape.

5.6 Allotropy

The existence of an element in more than one crystalline forms is known as allotropy. These forms of the element are called allotropes or allotropic forms. This happens when the atoms of the element are bonded together in a different way. Different bonding arrangements between atoms result in different structures with different physical properties. Only some elements like sulphur, phosphorus, carbon show allotropy.

For example, the allotropes of carbon include:

- Diamond, where the carbon atoms are bonded together in a four-cornered lattice arrangement.

Fig 5.18 Diamond

- Graphite, where the carbon atoms are bonded together in sheets of a six-sided lattice.

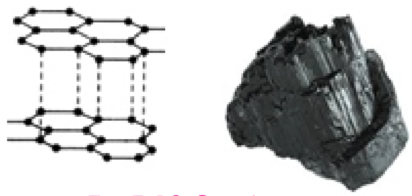


Fig 5.19 Graphite

- Graphene, single sheets of graphite.

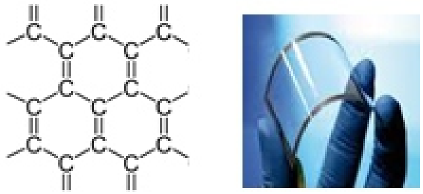


Fig 5.20 Graphene

- Fullerenes, where the carbon atoms are bonded together in spheres, cylinders or egg-shaped formations.

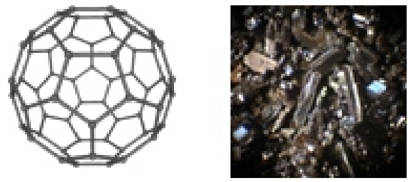


Fig 5.21 Fullerenes

5.7 Plasma State

English scientist William Crookes identify the fourth states of matter known as Plasma. It was discovered by adding energy to a gas. As a result some electrons left their atoms and formed positive and negative ions by ionization. In plasma these charged particles react strongly electric and magnetic fields. If plasma loses heat, the ions will re-form into a gas, emitting the energy which had caused them to ionize.

Fig 5.22

It means that plasma is distinct state of matter containing a significant number of electrically charged particles which affect its electrical properties and behavior.

Some examples of daily life are as follow:

- The lightning makes plasma naturally.
- The Artificial (man-made) uses of plasma include fluorescent light bulbs, Neon signs.
- The use of plasma display of television or computer screens.
- The plasma lamps and globes are popular in children's toys and room decoration.
- Scientists are experimenting with plasma to make a new kind of nuclear power, called fusion, which will be much better and safer than ordinary nuclear power with less radioactive waste.

Fig 5.23

5.8 Bose Einstein Condensate (BEC)

Two scientists Satyendra Bose and Albert Einstein discovered another state of matter in 1920 but they did not have the equipment and facilities to make it happens at that time. Afterward in 1995 two other scientists Eric Cornell and Carl Weiman also proposed another state of matter known as Bose-Einstein Condensate (BEC)

They discovered that as plasma are super-hot and super excited atoms. The atoms in a Bose-Einstein condensate (BEC) are totally opposite. They are super unexcited and super cold atoms.

Let's understand condensation first. Condensation happens when several gas molecules come together and form a liquid. It all happens because of a loss of energy. Gases are really excited or energetic atoms. When they lose energy, they slow down and come close to each other. They can gather into one drop.

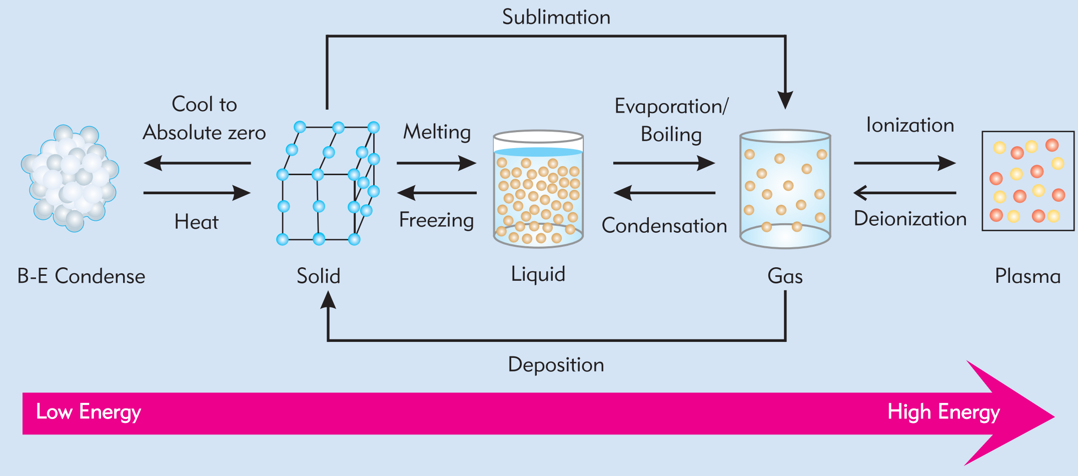
For example, when you boil water. Water vapour in the form of steam condenses on the lid of pot. Vapour cool and become a liquid again. You would then have a condensate.

Do you know?

Absolute Zero is a temperature where all the atomic activities stop theoretically.

The BEC happens at super low temperatures. When we get a temperature near absolute zero, all molecular motion stops due to depletion of energy, and atoms begin to clump. The result of this clumping is the BEC. In common we can not see this state in observable nature, because it is very difficult to reach at very low temperature under normal lab conditions.

Do you know?



Summary

- The states of matter are solid, liquid, gas, plasma, and BE condensate.
- The kinetic particle theory states that all matter is made up of tiny particles that are in constant random motion.
- Solids have fixed shapes and volumes. They cannot be compressed.
- Liquids do not have fixed shapes. They have fixed volumes and cannot be compressed easily.
- Gases do not have fixed shapes or volumes. They can be compressed.
- Boyle's law states that volume of a given mass of a gas is inversely proportional to the pressure at constant temperature.
- Charles law states that volume of a gas is directly proportional to the absolute temperature at a constant pressure.
- Diffusion is mixing up of a gas molecules throughout a space or other gases. Gases diffuse very rapidly.
- Effusion is escaping of a gas molecule through a fine hole into an evacuated space.
- The pressure exerted by vapours in equilibrium with its pure liquid at a particular temperature is called Vapour Pressure.
- The temperature at which vapour pressure of a liquid become equal to atmospheric pressure is called boiling point of liquid
- The Density of liquid depends upon mass per unit volume. Affected by temperature and pressure.
- The temperature at which a solid starts melting and coexist in equilibrium with liquid state is called melting point.
- Freezing point of a liquid is that temperature at which vapour pressure of a liquid state is equal to the vapour pressure of a solid state.
- Solids are rigid and denser than liquid and classified as amorphous and crystalline.
- Amorphous solids are shapeless and do not have sharp melting point.
- Crystalline solids have definite three dimensional pattern of arrangement of molecules.They have sharp melting point.
- The solids exist in different physical forms is known as allotropy.
- Plasma are super-hot and super excited atoms.
- The atoms in a Bose-Einstein condensate (BEC) are super unexcited and super cold atoms.

Exercise

SECTION- A: MULTIPLE CHOICE QUESTIONS

Tick Mark (✓) the correct answer

1. Which of the following gases diffuse fastest?
   (a) Hydrogen (b) Chlorine
   (c) Fluorine (d) Helium
2. The vapour pressure of a liquid increases with the:
   (a) increase of pressure
   (b) increase of temperature
   (c) increase of intermolecular forces
   (d) increase the polarity of molecules
3. The freezing point depends upon:
   (a) Nature of liquid (b) Pressure
   (c) Temperature (d) Volume
4. One atmospheric pressure is equal to:
   (a) 10325 Pascal (b) 106075 Pascal
   (c) 10523 Pascal (d) 101325 Pascal
5. Which of the following does not affect the boiling point :
   (a) Intermolecular forces (b) External pressure
   (c) Initial temperature of liquid (d) Nature of liquid
6. The mobility of liquids is lesser than:
   (a) Solids (b) Gases
   (c) Both a & b (d) None of these
7. Which of the following have sharp melting point in solids :
   (a) Plastic (b) Rubber
   (c) Glass (d) Diamond
8. Which of the following is the lightest form of matter:
   (a) Solid (b) Liquid
   (c) Gases (d) Plasma
9. The liquid molecules leave the surface of liquid in evaporation process because :
   (a) Energy is low (b) Energy is moderate
   (c) Energy is high (d) None of these
10. The density of gases increases when its :
   (a) Pressure increased (b) Temperature increased
   (c) Volume increased (d) None of these

SECTION- B: SHORT QUESTIONS:

1. Define the allotropy with examples?
2. What is Effusion, explain with the examples?
3. Define the following?
   Boiling Point, Melting Point, Freezing Point
4. What is density, how the density of liquid is affected by temperature and pressure?
5. Explain plasma with the daily life examples?
6. Justify that atoms of Bose Einstein condensate are super unexcited and super cooled?
7. How kinetic molecular theory differentiates states of matter?

SECTION- C: DETAILED QUESTIONS:

1. Discuss the property of evaporation in liquids? Which factors affects the evaporation process?
2. Describe the Boyl's law with example?
3. Differentiate between amorphous and Crystalline Solids?
4. Define and explain the Charles' law of gases?
5. Describe the process of diffusion in liquids? State the factors which influence it?
6. How boiling point is affected by different factors?
7. Define vapour pressure and justify that it is visible in a close system only?

SECTION- D: NUMERICALS

1. Convert the following units :
   (A)100°C to K (b) 150°C to K
   (c) 780K to °C (d) 170 K to °C
2. it is desired to increase the volume of a fixed amount of gas from 90.5 to 120 cm^3 while holding the pressure constant. What would be the final temperature if initial temperature is 33°C.
3. A 78ml sample of gas is heated from 35°C to 80°C at constant pressure. What is the final volume?
4. A gas occupies a volume of 40.0 dm^3 at standard temperature (0°C) and pressure (1 atm), when pressure is increases up to 3 atm at unchanged temperature what would be the new volume?
5. The 800 cm^3 of a gas is enclosed in a container under a pressure of 750 mm. If the volume is reduced to 250 cm^3, what will be the pressure?
6. The pressure of a sample gas is 8 atm and the volume is 15 liters. If the pressure is reduced to 6 atm, what is the volume?

Chapter 6

SOLUTIONS

Time Allocation

Teaching periods	= 15
Assessment period	= 3
Weightage	= 16

Major Concepts

6.1 Solution, aqueous solution, solute, and solvent

6.2 Saturated, unsaturated, supersaturated solutions and dilution of solution

6.3 Types of solution

6.4 Concentration Units

6.5 Solubility

6.6 Comparison of Solutions, Suspension, and Colloids

STUDENTS LEARNING OUT COMES (SLO'S)

By the end of Chapter Students will be able to:

- Define the terms: solution, aqueous solution, solute and solvent and give an example of each. (Remembering)
- Explain the difference between saturated, unsaturated and supersaturated solutions. (Analyzing)
- Explain the formation of solutions (mixing gases into gases, gases into liquids, gases into solids) and give an example of each. (Understanding)
- Explain the formation of solutions (mixing liquids into gases, liquids into liquids, liquids into solids) and give an example of each. (Understanding)
- Explain the formation of solutions (mixing solids into gases, solids into liquids, solids into solids) and give an example of each. (Understanding)
- Explain what is meant by the concentration of a solution. (Understanding)
- Define Molarity. (Remembering)
- Define percentage solution. (Remembering)
- Solve problems involving the Molarity of a solution. (Applying)
- Describe how to prepare a solution of given Molarity. (Applying)
- Describe how to prepare dilute solutions from concentrated solutions of known Molarity.
- (Applying)
- Conversion between the Molarity of a solution and its concentration in g/dm^3. (Applying)
- Use the rule that "like dissolves like" to predict the solubility of one substance in another. (Understanding)
- Define colloids and suspensions. (Remembering)Differentiate between solutions, suspension and colloids (Analyzing)

Introduction

A solution is the homogenous mixture of solute and solvent. Solutions are everywhere around us. Many substances around us are solution as like a glass of milk, drugs and medicines, blood, alloy, kerosene oil, tap water, cooking utensils and surgical tools. These all are the example of a solution. The nutrient absorbed by plants from the soil is also the example of the solution. The foods we eat come into contact enzymes with the help of a solution. The majority of chemical reactions occur in solutions. All these are possible with the presence and support of the solution.

In this chapter, we will study about solutions, types of solution and comparison of solutions, suspension and colloids in detail.

6.1 Solution, Aqueous Solution, Solute and Solvent

6.1.1 Solution

A solution is a homogeneous mixture of two or more substances to form a single phase. A solution exists in all three states of matter. An example of solid solution is brass (Zinc dissolved in copper), the liquid solution is sugar into water and gaseous solution is air we breathe. Air is made up of gases like oxygen, nitrogen, carbon dioxide etc.

6.1.2 Aqueous Solution

An aqueous solution is formed by dissolving a substance in water. The word aqueous is derived from the Latin word called aqua meaning water. Sugar, salt, and acid in water are the examples of aqueous solution. In aqueous solution, water (H_2O) is present in greater amount and termed as solvent.

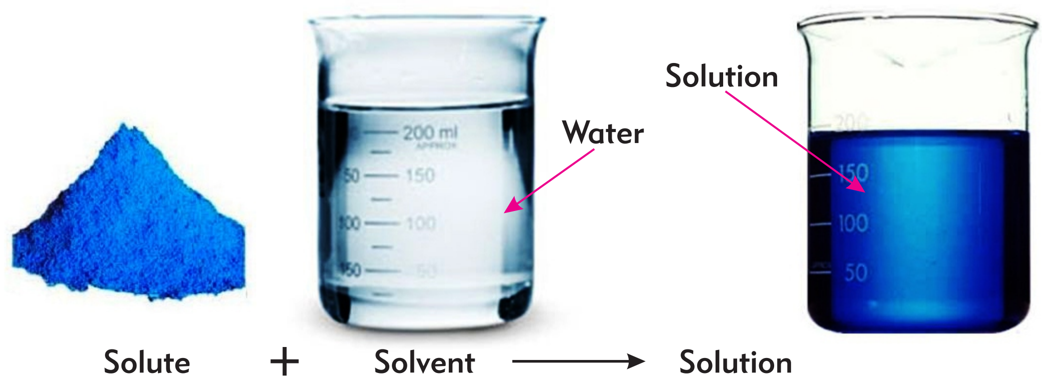


Figure 6.1 Preparation of Solution

6.1.3 Solute

The component of solution which is always present in smaller amount is called the **solute**. A solute is dissolved in a solvent to make a solution. An everyday example of a solute is sugar in water. Sugar is the solute and water is solvent. In a solution more than two solute may be present. For example, in a soft drink, sugar, salt and carbon dioxide are solute and water is a solvent. Consider another example, air is a solution of several gases like nitrogen, carbon dioxide, oxygen and inert gases. In this solution carbon dioxide, oxygen and inert gases are solute and nitrogen is solvent.

6.1.4 Solvent

The component of the solution which is present in larger amount is called **solvent**. Solvent are generally liquid, but can also be gas or solid. The component of solution which can dissolve solute is called solvent. Water is the most common solvent because it can dissolve most of the solutes. It is also called as the **universal solvent.**

Test Yourself

- Why solutions are important for us?
- Why solution is called mixture?
- How is solution formed?
- What is an aqueous solution?
- Write any two examples of solute and solvent.
- Air is solution of general gases like nitrogen, carbon dioxide, oxygen and inert gases, why nitrogen is called solvent?

6.2 Saturated, Unsaturated, Supersaturated Solutions and Dilution of the Solution

6.2.1 Saturated solution

Take some water in a beaker and add sugar in a small amount. The sugar dissolves very easily in the water at a certain temperature. If the adding of sugar continues, it is found that a limit is reached when the water cannot dissolve any more sugar and settling down occurs and no more solute can dissolve in it. Such a solution is said to be saturated solution. Thus we can define a saturated solution as follows:

A solution which cannot dissolve more solute in it at a particular temperature is called a **saturated solution.**

6.2.2 Unsaturated solution

A solution which contains lesser amount of solute than is required to saturate it at a particular temperature, is called **unsaturated solution.** Salt dissolved in water is the example of unsaturated solution if the solution has ability to dissolve more solute to become a saturated solution.

6.2.3 Supersaturated solution

When we heat the saturated solution, it develops further capacity to dissolve more amount of sugar (solute). This solution contains greater amount of solute than that present in a saturated solution and the solution become more concentrated. Now solution is called super saturated.

Thus, we can define a supersaturated solution as:

A solution that can dissolve more solute than it contained in the saturated solution after heating is called a **Supersaturated solution.**

For better understanding of saturated, unsaturated and supersaturated solution, a brief comparison of their differences is given in table.6.1

Table 6.1
The difference between Saturated, Unsaturated and Supersaturated solution

Saturated	Unsaturated	Supersaturated
In saturated solution maximum amount of solute that can be dissolved at particular temperature.	In unsaturated solution less amount of solute is dissolved at particular temperature.	In super saturated solution more amount of solute has been dissolved than its maximum capacity.
The solution has high concentration than unsaturated solution.	The solution has low concentration than saturated solution.	The solution has more concentration than saturated solution.
There is no formation of precipitation at the bottom of container.	There is also no precipitation at the bottom of container.	There is formation of precipitation at the bottom of container on cooling.
A solution having 36 gram of sodium chloride per $100cm^3$ of water at 20°C is the example of saturated solutions.	A solution having amount of salt less than 36 gram per $100cm^3$ of water at 20°C is the example of unsaturated solution.	A solution having more amount than 36 gram of salt per $100cm^3$ of water at higher temperature is the example of supersaturated solution.

6.2.4 Dilution of solution

We already know about two basic terms dilute and concentrated solutions and both are depending on the relative amount of solute present in it. **Dilute solution** contains a relatively small amount of a solute in a large amount of solvent like adding more water to a solution.

Whereas, **concentrated solution** contains a relatively large amount of solute in a small amount of solvent.

Dilution of a solution is necessary process in the laboratory, as stock solution (more concentrated) are often purchased and stored in laboratory when the desired concentration of solution is required can be prepared by diluting the stock solution by given formula.

Preparing dilute solution

In a laboratory, we can make a dilute solution from a concentrated solution by using formulas:

Concentrated Solution = Dilute Solution

$$M_1V_1 = M_2V_2$$

Where, M_1 = Molarity of concentrated solution

V_1 = Volume of Concentrated solution

M_2 = Molarity of dilute solution

V_2 = Volume of dilute solution

Note: $1 cm^3 = 1 ml$ and $1 dm^3 = 1 L$

6.2 Dilute solution

Example 6.1

How do you prepare solution of 100ml of 0.40M $MgSO_4$ from a stock solution of 2.0M $MgSO_4$?

Solution:

Data M_1 = 2.0M $MgSO_4$

M_2 = 0.40M $MgSO_4$

V_2 = 100ml

V_1 = ?

Concentrated Solution = Dilute Solution

$$M_1 V_1 = M_2V_2$$

$$2 \times V_1 = 0.40 \times 100$$

$$V_1 = \frac{0.4 \times 100}{2}$$

$$= 20 \text{ cm}^3$$

Transfer 20cm^3 of 2 M $MgSO_4$ to a 100cm^3 measuring flask and dilute it by adding water up to the marks. It is 0.40M solution of $MgSO_4$

Example 6.2

How would you prepare 500 cm^3 of 0.20 M NaOH (aq) from a stock solution of 1.5 M NaOH?

Solution

M_1 = 1.5M NaOH
M_2 = 0.2M NaOH
V_2 = 500cm^3
V_1 = ?

Concentrated Solution = Dilute Solultion

$$M_1V_1 = M_2V_2$$

$$1.5 \times V_1 = 0.20 \times 500$$

$$V_1 = \frac{0.20 \times 500}{1.5}$$

$$= 66.67\text{cm}^3$$

Take 66.67 cm^3 of concentrated NaOH and placed in a 500cm^3 measuring flask and dilute it by adding water up to the mark. It will become 0.2M solution of NaOH.

Test Yourself

- Consider two beakers A and B. Each one having 20ml of water. Add 10 g of sodium thiosulphate in beaker A and 20 g in beaker B and stir carefully. Answer the following questions:
- Which beaker's solution is saturated?
- How can be supersaturated solution prepared from above experiment?
- Compare the term saturated and unsaturated solution.
- Which beaker's solution is unsaturated solution?
- How is unsaturated solution prepared?
- 10M HNO_3 solution is available in laboratory. How would you prepare 500 cm^3 of 0.1M solution?

6.3 TYPES OF SOLUTION

As we know that there are three states of matter, i.e. solid, liquid and gas. A solute as well as solvent can exist in any one of the three states of matter. By the combination of these three states different types of solutions are formed, which are given in table 6.2

Table 6.2 Types of Solution with Examples

Sr. No	State of Solute	State of Solvent	State of Solution	Example of Solution
1.	Gas	Gas	Gas	Air, mixture of hydrogen and helium in water balloon, oxygen in air
2.	Gas	Liquid	Liquid	Carbonated drinks (carbon dioxide dissolved in water)
3.	Gas	Solid	Solid	Hydrogen gas in palladium, nitrogen in titanium.
4.	Liquid	Gas	Gas	Fog (water vapour in air) liquid air pollutants
5.	Liquid	Liquid	Liquid	Alcohol dissolves in water , oil in ether.
6.	Liquid	Solid	Solid	Amalgam, butter, cheese
7.	Solid	Gas	Gas	Smoke (carbon particles in air)
8.	Solid	Liquid	Liquid	Salt in water, Sugar in water
9.	Solid	Solid	Solid	Brass an alloy (Zinc dissolved in copper), Bronze (tin dissolved in copper)

6.4 CONCENTRATION UNITS

We already discussed in section 6.2.4 that concentration is the amount of solute present in a given amount of solvent or solution. It is also the ratio of amount of solute to the amount of solution or ratio of amount of solute to the amount of solvent.

We can express the concentration in terms of grams of solute per dm^3 (g/dm^3) of solution.

$$\text{Concentration in } g/dm^3 = \frac{\text{Mass of solute in gram}}{\text{volume of solution in } dm^3}$$

Do you know?

$1\ L = 1 dm^3$

$1 ml = 1 cm^3$

$1 L = 1000 ml$

$1 dm^3 = 1000 cm^3$

There are many ways of expressing the concentration of the solution. In this class we will discuss only two main concentration units.

6.4.1 Percentage

It is a unit of concentration. It refers to the percentage of solute present in a solution. It can be expressed in four different ways:

(i) Mass by mass percent (%m/m)

It is the mass of solute in gram dissolved in 100 grams of solution.

For example, 5% m/m sugar solution means that 5 grams of sugar dissolved in 95 grams of water to make 100 grams of solution.

$$\text{Percent solution } \left(\frac{m}{m}\%\right) = \frac{\text{Mass of Solute (g)}}{(\text{Mass of Solute + masss of solvent (g)})} \times 100$$

or

$$= \frac{\text{Mass of solute (g)}}{\text{Mass of solution (g)}} \times 100$$

(ii) Mass by volume percent (%m/v)

It is the mass of the solute in grams dissolved per $100 cm^3$ of the solution.

For example, 5%m/v sugar solution means that 5 grams of sugar in $100 cm^3$ of the solution.

(iii) Volume by mass percent (%v/m)

It is the volume of solute in cm^3 dissolved in 100 gram of the solution.

For example, 5%(v/m) Alcohol solution means $5cm^3$ of alcohol is dissolved in (unknown) volume of water so that mass of solution become 100g.

$$\text{Percent of solution } \frac{V}{m} = \frac{\text{Volume of Solute } (cm^3)}{\text{Mass of Solution (g)}} \times 100$$

(iv) Volume by volume percent (%v/v)

The volume of solute in cm^3 is dissolved per $100cm^3$ of the solution.

For example, 5%(v/v) Alcohol solution means that $5cm^3$ of alcohol is dissolved in 95 cm^3 of the water to make 100 cm^3 of the solution.

$$\text{Percent of solution } \frac{V}{V} = \frac{\text{Volume of Solute } (cm^3)}{\text{Mass of Solution } (cm^3)} \times 100$$

Example 6.3 (Percent by Mass)

Calculate the percentage concentration (m/m) of the solution obtained by dissolving 15g salt in 110g water.

Solution

Mass of salt = 15 g

Mass of water = 110 g

Mass % of salt =?

Total mass of solution = 15 g salt + 110 g water = 125 g

Percent by mass would be calculated as:

$$\text{Percent (mass/mass)} = \frac{\text{Mass of Solute (g)}}{\text{Mass of Solution (g)}} \times 100$$

$$= 15/125 \times 100 = 12\%$$

Thus the concentration of a solution is 12% by mass.

Example 6.4 (Percent by Volume)

Calculate the volume/volume percent of solution obtained by mixing $25cm^3$ of ethanol in water to produce $150cm^3$ of the solution.

Solution

Volume of solute = $25cm^3$

Volume of solution = $150cm^3$

Volume/volume percent =?

$$\text{Percent of solution } \frac{V}{V} = \frac{\text{Volume of Solute } (cm^3)}{\text{Volume of Solution } (cm^3)} \times 100$$

$$= \frac{25}{150} \times 100$$

$$= 16.7\%$$

Thus the concentration of a solution is 16.7% by volume

6.4.2 Molarity

Molarity is defined as the number of moles of solute dissolved in one dm^3 (1 Litre)of the solution. Molarity is the concentration unit in which the amount of solute is expressed in gram and the volume of solution is expressed in dm^3. It is denoted by "M" and its unit is mol/dm^3.

$$\text{Molarity (M)} = \frac{\text{Number of moles of solute}}{\text{Volume of solution in } dm^3}$$

$$\text{Number of moles of solute} = \frac{\text{Mass of solute}}{\text{Molar mass of the solute } (gmol^{-1})}$$

$$\text{Volume of solution in } dm^3 = \frac{\text{Volume of solution } (cm^3)}{1000}$$

$$\therefore \quad \text{Molarity} = \frac{\text{Mass of solute (g)}}{\text{Molar mass of the solute } (gmol^{-1})} \times \frac{1000}{\text{Volume of solution } (cm^3)}$$

Preparation of Molar Solution

One mole (molar mass) of a solute is dissolved in a sufficient amount of water so that the total volume is $1dm^3$. The solution is said to be one molar solution.

For example, prepare 1.0 M solution of NaCl in 1 dm^3.

Following steps may be considered:

1. Weigh out 58.5 grams of NaCl — Molar mass of NaCl = 23 + 35.5 = 58.5g/mol
2. Put the NaCl into a volumetric flask
3. Add water to dissolve the salt and prepare 1 dm^3 solution.

You have prepared 1M NaCl solution by dissolving 58.5 grams of NaCl in sufficient water to make the volume 1 dm^3.
Similarly, to make a 0.1M solution, you could dissolve 5.85g of NaCl in 1 dm^3 of water.

Problems based on Molarity of a solution

Example 6.5

20 gram of salt (NaCl) is dissolved in 500 cm^3 of a solution. Calculate the molarity of that solution.

Data:

$$\begin{aligned} \text{Mass of solute} &= 20g \\ \text{Molar mass of NaCl} &= 23 + 35.5 \\ &= 58.5g/mol \\ \text{Volume of solution} &= 500cm^3. \\ \text{Molarity (M)} &= ? \end{aligned}$$

Formula

$$\text{Molarity} = \frac{\text{Mass of solute (g)}}{\text{Molar mass of the solute (gmol}^{-1})} \times \frac{1000}{\text{Volume of solution (cm}^3)}$$

Solution:

$$\text{Molarity} = \frac{20g}{58.5g/mol} \times \frac{1000}{500}$$

$$\text{Molarity} = 0.683 \text{ mole / dm}^3.$$

Example 6.6

What is the mass of oxalic acid present in 100 cm^3 of 2 molar solution?

Data:

Molarity = 2 mol/dm^3

Volume in cm^3 = 100

Molar mass of oxalic acid ($C_2H_2O_4$) $= (12 \times 2) + (1 \times 2) + (16 \times 4)$

$= 24+2+64 = 90 g/mol$

Mass of solute = ?

Formula

$$\text{Molarity} = \frac{\text{Mass of solute (g)}}{\text{Molar mass of the solute (gmol}^{-1})} \times \frac{1000}{\text{Volume of solution (cm}^3)}$$

Solution:

$$2 = \frac{\text{Mass of solute}}{90} \times \frac{1000}{100}$$

$$\text{Mass of solute} = \frac{2 \times 90 \times 100}{1000} = 18gm$$

Example 6.7

A sample of sulphuric acid has the molarity 20M. How many cm^3 of solution should we use to prepare 500 cm^3 of 0.5M H_2SO_4?

Solution:

M_1 = Molarity of given H_2SO_4 = 20M

M_2 = Molarity of required H_2SO_4 = 0.5M

V_2 = Volume required in H_2SO_4 on = 500 cm^3

V_1 = volume of concentrated solution needs for dilution =?

By using dilution formula

$M_1V_1 = M_2V_2$

$$V_1 = \frac{M_2V_2}{M_1} = \frac{0.5 \times 500}{20}$$

V_1 = 12.5 cm^3

12.5cm^3 of 20M is used to make 500cm^3 aqueous solution to form 0.5M H_2SO_4.

Test Yourself

- Define concentration
- Brass contains 20% zinc and 80% copper. Identify state of solute and solvent in this solution? Also write the type of solution.
- Write the difference between diluted and concentrated solution.
- Which one of the solution is more diluted; 2M or 3M?
- A solution of NaOH has concentration 1.2M, calculate the mass of NaOH in g/dm^3 in this solution (50 cm^3).
- Determine the percentage concentration of the solution obtained by dissolving 10g sugar in 140g water
- A student asked to prepare 10%(m/m) solution of sugar. How much solvent will be required to prepare this solution?

6.4 SOLUBILITY

Solubility is defined as the maximum quantity of solute that can be dissolved in 100 grams of solvent to prepare saturated solution at a particular temperature.

Different substances have different capability to dissolve in the same solvent at a particular temperature. For example, the solubility of sodium chloride in 100g of water at 100°C is 39.12g, whereas, the solubility of silver chloride in 100g of water at 100°C is 0.002g. This shows that sodium chloride is much more soluble in water than silver chloride.

General Principles of Solubility

1. The general principle of solubility is **"Like dissolves like"**. Its mean that two substances with similar intermolecular forces are likely to be soluble in one another. It has been observed that:
   - Ionic and polar solute dissolved in polar solvents. For example, Na_2CO_3, sugar and alcohol are polar and dissolved in water because water is also polar.
   - Nonpolar solute dissolved in non-polar solvents. Such as oil and paints are non-polar, they are dissolved in ether as both are non-polar. Similarly, waxes and fats dissolve in benzene and not in water.
   - Nonpolar compounds are not soluble in polar solvents (water). For example, oil, petrol, benzene are non-polar, they are not dissolved in water because water is polar.
2. Solute-solvent interaction
3. Temperature

Do you know?

When solute-solute or solvent-solvent interactions are much more than solute-solvent interaction, a solution will not form.

6.5.1 Solubility and Solute-Solvent Interaction

In order to dissolve the solute into solvent, the following conditions must be fulfilled.

- Solute-Solute bonding should be broken.
- Solvent-Solvent bonding should be broken to provide space for solute particles.
- Solute-Solvent attraction should be maximized.

The process of solution formation depends upon the relative strength of attractive forces between solute-solute, solvent-solvent and solute-solvent. A solute will dissolve in a solvent if the solute-solvent forces of attraction are greater enough to overcome the solute-solute and solvent-solvent forces of attraction. A solute will not dissolve if the solute-solvent forces of attraction are weaker than individual solute and solvent intermolecular attractions.

As we know that sodium chloride is an ionic compound. When sodium chloride (NaCl) is added into water, it dissolves quickly. The negative end of water molecules is attracted to sodium ions and the positive end of water molecules is attracted to chlorine ions. In this case, solute-solvent attractions are more in comparison with solute-solute interaction, therefore solution of sodium chloride is formed. These attractive forces of water are stronger enough to overcome the attraction between Na^+ and Cl^- ions in NaCl. Figure 6.3 shows the attraction of Na^+ and Cl^- ions with water molecules.

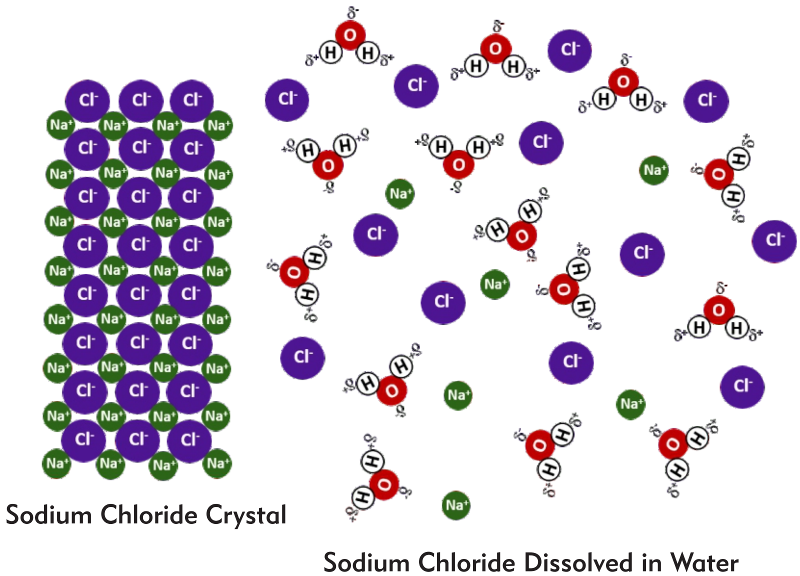


Figure 6.3 Interaction of sodium chloride and water

Now it is concluded that if there is more attraction between solute and solvent in comparison with solute-solute interaction, then the solution will be formed. If Solute-solute interaction is greater than solute-solvent attraction, then solute will not dissolve in solvent.

6.5.2 Effect of Temperature on Solubility

Solubility is directly proportional to the temperature for solid & liquid solutes. Solubility is increased by increasing the temperature because hot water molecules have greater kinetic energy and collide with solid solute more vigorously. For example, a greater amount of sugar will dissolve in warm water than in cold water. The solubility of potassium chloride is 277.7g/L in water at 20°C. It will become 540.2g/L at 100 °C.

For all gases, the solubility decreases as the temperature of the solution increases.

Test Yourself

- Explain the general principle for solubility, "like dissolves like"?
- Why solute dissolves in solvent?
- Why benzene does not dissolve in water?
- Suppose solute-solute forces are weaker than those of solute-solvent forces. Will solute dissolve in solvent and to be form a solution?
- What is the main reason that non-polar solute does not dissolve in water?

6.6 COMPARISON OF SOLUTION, SUSPENSION AND COLLOIDS

When a solute (sugar or table salt) is placed in water, after some time sugar or salt completely dissolve in water and we cannot even see the particles of sugar or salt. Now if you repeat the same practice with sand or clay, would you get same results? The solution of sugar in water is a clear solution, whereas sand or clay in water is not a clear solution. After some time, sand or clay settles at the bottom and we can see the particles of sand clearly. Now compare these two solutions with milk. Milk is not a clear solution but particles do not settle at the bottom with time. So we can say that particles remain dispersed in solution, but the size is big enough that does not allow the clear appearance of the solution.

On the basis of particles size and their characteristics, mixtures can be classified as true solution, colloid and suspension.

6.6.1 Solution

A solution is a homogeneous mixture of two or more than two components. When we dissolve sugar in water, after some time sugar completely dissolve in water and we cannot even see the particles of sugar. Sugar is dissolved into the water and a drop of ink mixed in water are the examples of the true solution. The solute particles in true solutions are very tiny so that we cannot see them with naked eyes.

6.6.2 Colloid

In a colloid, the particles are larger than those present in a true solution, but smaller than the particles that make up a suspension. Therefore, they will not settle to the bottom of the container and remain dispersed in the medium. The dispersed particles of colloids cannot be separated by filtration, but they scatter the beam of light. This phenomenon is called the Tyndall effect. These solutions are also called false solutions. Milk, butter, jelly, blood are the example of colloidal solution. Few other examples of colloid are fog, smoke and dust particles which are suspended in the air and the beam of light can be easily scattered.

Do You Know?

The scattering of visible light by Colloidal particles is called Tyndall effect. This phenomenon was discovered by Physicist John Tyndall in the 19th century.

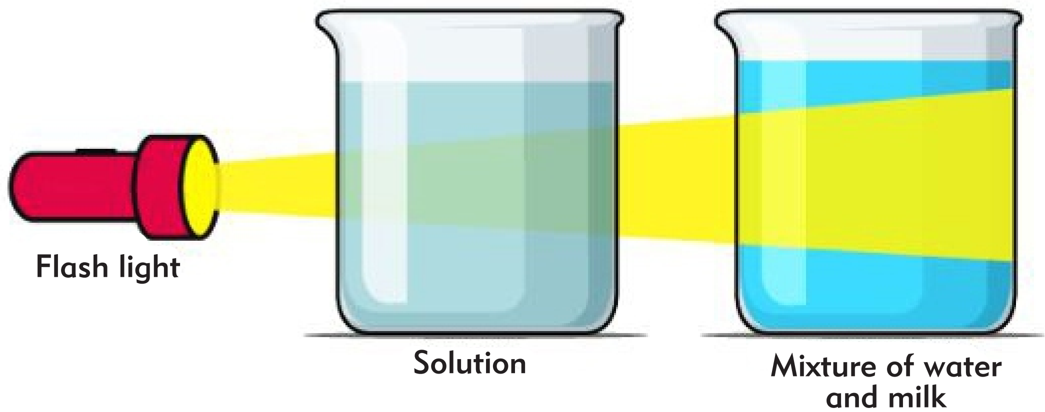


Figure: 6.4Tyndall effect by colloids

6.6.3 Suspension

A suspension is a heterogeneous mixture of solute and solvent in which solute particles do not dissolve. A suspension contains large particles suspended in a liquid. The particles will eventually, slowly settle to the bottom in the absence of agitation. For examples mud in water, chalk in water, paints etc. The solute particles are big enough to be seen with naked eyes and they can scatter light like colloids.

Table 6.3 shows the comparison of the characteristics of solution, suspension and colloid.

Table 6.3 Comparison of the characteristics of solution, Suspension and Colloid.

Solution	Suspension	Colloid
Particle size less than 1 nm	Particle size greater than1000 nm	Particle size 1 to 100 nm
Homogeneous (particles dissolve uniformly)	Heterogeneous (particles settle down after sometimes)	Homogeneous and heterogenous (Particles do not settle down for a long time)
Particles cannot be distinctly seen with the naked eye.	Particles are big enough but can be seen with naked eyes	Colloidal particles can not be with the naked eye but can be seen through ultra microscope
Clear, transparent and homogeneous	Cloudy, non uniform and heterogeneous	Cloudy, homogeneous, and uniform
Transparent but often colored	Translucent and often opaque	Often opaque, but can be transparent
Solutes cannot easily be separated	Solutes can be separated easily	Solutes can not be separated easily
Do not scatter light	Scatter light, but are not transparent	Scatter light (Tyndall effect)
Particles can pass through filter paper	Particles do not pass through filter paper	Particles pass through filter paper

Test Yourself

- How is a colloid different from a solution?
- Which one is colloidal solution, Starch solution or Glucose solution?
- Paint is colloidal solution. Give reason?
- Why paints should be stirred thoroughly before using?
- Write two examples of suspension.
- Differentiate between suspension and colloid.
- Justify that milk is a colloidal solution.
- Why sugar solution don't scatters the light?
- Define false solution.
- Why does the colloidal show tyndall effect?

Society, Technology and Science

Relate Solution to Different Products in the Community

Solutions have great importance and influence in our daily life. As we look in our surrounding, such as air, soft drinks, beverages, medicines, butter, tooth pastes, natural gas and even water are solutions. When we stirr sugar in a cup of tea, we are making a solution. Most of chemical reactions take place in the bodies of living organism occur in presence of water as a solvents. The food assimilation process in our bodies also occurs in solution. Brass and steel are also solution. These solutions are widely used for making cooking utensils, surgical instruments, cutlery, and many other objects. Silver and tin amalgams are widely used to make dental filling. The majority of chemical processes are reactions that occur in solution. Gaseous solutions are also used in chemical industries for the preparation of urea, ammonia gas, nitric acid, rubber, vegetable oil etc.

Summary

- A **solution** is a **homogeneous mixture** of two or more substances.
- The substance which is dissolved is called **solute**.
- The substance in which the solute is dissolved is called a **solvent.**
- The solution is single phase homogeneous mixture of solute and solvent.
- Component of solution which is present in small quantity and it can be dissolved in solvent is called **"solute"**.
- Component of solution which is present in large quantity and it can dissolve solute is called **"solvent"**.
- The aqueous solution is a type of solution, in which water used as a solvent.
- In unsaturated solution the amount of solute is less than its original capacity to dissolve.
- In a saturated solution maximum amount of solute dissolves in solvent according to their capacity.
- In supersaturated solution the capacity to dissolve the solute is increased by increasing the temperature.
- A solution has nine types based on the nature of solute and solvent. A solute may be solid, liquid and gas. If the solution is in liquid form so it is considered a true solution.
- A dilute solution has a small amount of solute in a large amount of solvent.
- A concentrated solution has a large amount of solute in small amount of solvent.

- We can dilute the solution by the formula. $M_1V_1=M_2V_2$.
- The proportion of solute in a solution is called concentration.
- Molarity is defined as the number of moles dissolved in a dm^3 of the solution. Those solutions whose concentration is expressed in molarity are called molar solution.
- Percent solution is based on mass and volume of the components of solute and solvent.
- The quantity of solute and solvent may be increase or decrease according to the percentage of the solution. Percent solution has four types
- Solubility is defined as the amount of solute dissolved in 100g of solvent. The main factors which effect the solubility are temperature, pressure and nature of solute and solvent.
- Nature of solute and solvent obey the principle of" like dissolves like".
- A heterogeneous mixture containing undissolved particles large enough to be seen with naked eyes is called suspension.
- In colloidal solutions, solute particles are larger than those present in the true solution but not large enough to be seen by naked eyes. They are also called false solution.

EXERCISE

SECTION- A: MULTIPLE CHOICE QUESTIONS

Tick Mark (✓) the correct answer

1. An alloy is the homogeneous mixture of:
   (a) two solids (b) two liquids
   (c) two gases (d) solid and liquid
2. A saturated solution of KCl on heating becomes:
   (a) unsaturated (b) supersaturated
   (c) diluted (d) all of these
3. If we dissolve sand into the water, then the mixture is said to be:
   (a) solution (b) suspension
   (c) colloids (d) concentrated solution
4. Solubility is usually expressed in grams of the solute dissolved in ________grams of a solvent
   (a) 10 (b) 100
   (b) 500 (d) 1000
5. Example of heterogeneous mixture is:
   (a) sugar and water (b) sand and water
   (c) salt and water (d) ink and water
6. 2 moles of sodium chloride (NaCl) is equal to:
   (a) 123 grams (b) 135 grams
   (c) 158 grams (d) 117 grams
7. Molarity of a solution which is prepared by dissolving 40g sodium chloride in 500cm^3 of solution is:
   (a) 1.4M (b) 1.5M
   (c) 1.6M (d) 1.7M
8. 10% (w/w) sugar solution mean that 10 grams of solute dissolved in:
   (a) 90g of water (b) 95g of water
   (c) 100g of water (d) 105g of water

9. An example of true solution is:
   (a) solution of starch (b) solution of soap
   (c) ink in water (d) tooth paste
10. Which solution contain more water:
   (a) 1.0M (b) 0.75M
   (c) 0.5M (d) 0.25M
11. When a saturated solution is diluted, it change into:
   (a) saturated solution (b) unsaturated solution
   (c) concentrated solution (d) supersaturated solution
12. Butter is example of solution:
   (a) gas-liquid (b) solid-liquid
   (c) liquid-solid (d) gas-solid
13. A solution that contain solid solute into liquid solvent is called:
   (a) solids in gas (b) liquids in solids
   (c) solids in solids (d) solids in liquid
14. What is the particle size in suspension?
   (a) 10^3nm (b) 10^2 nm
   (c) less than 10^3nm (d) greater than 10^3nm
15. Write the example of each type of solution.

Solute	Solvent	Example
Solid	Liquid	
Gas	Gas	
Solid	Solid	
Liquid	Solid	
Liquid	Gas	
Liquid	Liquid	

Section B: Short Questions

1. Explain the solute-solvent interaction to prepare sodium chloride solution.
2. Differentiate between the saturated and unsaturated solution?
3. Define solution and explain the major components of solution?
4. What do you mean by mass/volume percent ?
5. Define with example one molar solution.
6. Why does colloidal show Tyndall effect?
7. Define the terms
   i. Dilution ii. Concentration iii. Solubility iv. Molarity
8. Polar and ionic solutes dissolve in polar solvent only. Why?
9. Why does Polar solute not dissolve in nonpolar solvent?
10. How are solutions beneficial for community?
11. Why salt dissolves in water?
12. Air is a solution containing oxygen, carbon dioxide, nitrogen and other gases. Which one of the gas is called solvent and why?
13. Why gasoline does not dissolve in water?

Section C: Detailed Questions

1. Describe how to prepare dilute solution from concentrated solution.
2. Define the term solubility. How does nature of solute and solvent determine the extent of dissolution?
3. Why the solubility of a salt increases with the increase in temperature?
4. Explain the attraction of Na^+ and Cl^- ions for water molecule.
5. Explain the solubility with reference to "like dissolve like" principle.
6. What is the difference between solution, colloids and suspension?

Section D: Numerical

1. What is the molarity of the solution prepared by dissolving 1.25 g of HCl gas into enough water to make 30 cm^3 of solution?
2. A solution of potassium chloride was prepared by dissolving 2.5 g of potassium chloride (KCl) in water and making the volume up to 100 cm^3. Find the concentration of solution in mol/ dm^3.
3. A flask contains 0.25 M NaOH solution. What mass of NaOH is present per dm^3 of solution?
4. What volume of 0.5M acid is needed to neutralize 200ml of 4M base?
5. A mineral water bottle contains 28 mg of calcium in 100 cm^3 of solution. What is the concentration in g/dm^3?
6. A solution of 20cm^3 of alcohol is dissolved in 80cm^3 of water. Calculate the concentration (v/v) of this solution.
7. How much sodium hydroxide (NaOH) is required to prepare 400 cm^3 of 0.3M solution?

Chapter 7

ELECTROCHEMISTRY

Time Allocation

Teaching periods	= 12
Assessment period	= 3
Weightage	= 12

Major Concepts

7.1 Oxidation and Reduction

7.2 Electrochemical Cells

7.3 Corrosion and its Prevention

7.4 Alloy Formation

STUDENTS LEARNING OUT COMES (SLO'S)

Students will be able to:

- Define oxidation and reduction in terms of loss or gain of oxygen or hydrogen. .
- Explain oxidation and reduction in terms of loss or gain of electrons.
- Describe the nature of electrochemical processes.
- Sketch an electrolytic cell, label the cathode and the anode.
- Identify the direction of movement of cations and anions towards respective electrodes. .
- List the possible uses of an electrolytic cell.
- Sketch a Danniell cell, labeling the cathode, the anode, and the direction of flow of the electrons.
- Distinguish between electrolytic and Galvanic cells.
- Define corrosion.
- Describe rusting of iron as an example of corrosion.
- Summarize the methods used to prevent corrosion.
- Explain electroplating of metals on steel (using examples of zinc, Tin and chromium plating).
- Describe how a battery produces electrical energy.

INTRODUCTION:

In our daily life we use digital watches, calculators, cars and mobile phones powered by batteries or dry cells.

Extraction of metals like aluminum, copper and electro plating of metals are few applications of electrochemistry. It can be defined as the branch of chemistry which deals with electro chemical reactions, electrolyte and electrochemical cells is called electrochemistry.

"or" It deals with the conversion of electrical energy into chemical energy and chemical energy into electrical energy.

7.1 OXIDATION AND REDUCTION REACTIONS:-

The chemical reactions in which chemical energy changes into electrical energy or Vice Versa are called electrochemical reactions.

Oxidation may involve introduction of oxygen or removal of Hydrogen from a chemical substance.

Example:

$$C + O_2 \longrightarrow CO_2 \text{ (burning of coal)}$$

$$2C_8H_{18} + 25\,O_2 \longrightarrow 16\,CO_2 + 18\,H_2O \text{ (combustion of octane)}$$

$$N_2H_4 + O_2 \longrightarrow N_2 + 2H_2O \text{ (removal of Hydrogen from Hydrazine)}$$

Similarly, reduction may involve addition of Hydrogen or removal of oxygen from a chemical substance.

Example:

$$CuO + C \longrightarrow Cu + CO \quad \text{(reduction of copper oxide)}$$

$$CH_2 = CH_2 + H_2 \xrightarrow[(200^\circ C)]{Ni} CH_3 - CH_3 \quad \text{(Reduction of ethene)}$$

$$Fe_2O_3 + 2Al \longrightarrow Al_2O_3 + 2Fe + Heat \quad \text{(Reduction of Iron oxide)}$$

Oxidation and reduction reactions are electrochemical reactions. In electrochemistry oxidation and reduction reactions involve transfer of electrons.

The electrochemical reaction in which atom molecule or ion loses electron and its oxidation number increases is called oxidation reaction.

Example: $Cu \longrightarrow Cu^{++} + 2\bar{e}$ (Oxidation)

The electrochemical reaction in which atoms, molecule or ion accepts electron and its oxidation number decreases is called reduction reaction.

Example: $S + 2\bar{e} \longrightarrow S^{2-}$ (Reduction reaction)

Oxidation and reduction reactions can be summarized as

Table 7.1

Oxidation	Reduction
Addition of Oxygen	Addition of Hydrogen
Removal of Hydrogen	Removal of Oxygen
Loss of electron by a substance	Gain of electrons by a substance
Increase in oxidation number of a substance	Decreases in oxidation number of a substance

Oxidizing and Reducing Agents:-

Oxidation occurs due to oxidizing agent and reducing agent is responsible for reduction. Oxidizing agents are substances that accept electrons. Similarly reducing agents are substances which lose electrons.

Do You Know?

Photosynthesis is an example of redox reaction.

Photosynthesis produces glucose.

$$6CO_2 + 6H_2O + \text{sun light} \longrightarrow C_6H_{12}O_6 + 6O_2$$

Water (H_2O) under goes oxidation by losing electrons and form hydrogen ions. Carbon dioxide (CO_2) accept electrons and hydrogen ions to form glucose ($C_6H_{12}O_6$)

Example: $Zn + Cl_2 \longrightarrow ZnCl_2$

Zinc undergoes oxidation by losing electrons and it act as reducing agent while chlorine undergoes reduction by accepting electrons and act as oxidizing agent. A table of oxidizing agents and reducing agents is given below.

Table 7.2

Oxidizing Agent	Reducing Agent
H_2SO_4, HNO_3, $KMnO_4$, $K_2Cr_2O_7$, Cl_2, Br_2, I_2 etc.	Alkali metals, Al, H_2S, Zn, NaH, KH etc.

Test Yourself

- Identify the oxidizing and reducing agents from the following.
  1. Al 2. Na 3. H_2S 4. H_2SO_4 5. $KMnO_4$ 6. Zn
- Identify the oxidizing and reducing agents for the following reaction.
  $$H_2 + Cl_2 \longrightarrow 2HCl$$

7.2 ELECTROCHEMICAL CELLS:

The device which convert chemical energy into electrical energy or vise versa using redox reaction are called electrochemical cells.

The electrochemical reactions are carried out in electrochemical cells. It consists of two electrodes at which redox reaction occurs. The electrode at which oxidation takes place is called Anode and electrode at which reduction occurs is called Cathode. The reactions occurs at each electrode is called half cell reaction. The overall cell reaction is the combination of two half cell reactions. Each electrode is in contact with battery and electrolyte present in cell. Electrochemical cells are of two types.

(1) Electrolytic Cells **(2) Galvanic Cells or Voltaic Cells**

7.2.1 Concepts of Electrolyte:

An electrolyte consists of free moving ions and conduct electricity. Acids, bases and salts in molten or in aqueous solution form are electrolytes. Some strong and weak electrolytes are shown below in table 7.3.

Table 7.3

	Strong Electrolyte	Weak Electrolyte
Acids	HCl, HNO_3, HI, H_2SO_4	H_2S, H_2CO_3, CH_3COOH
Bases	KOH, NaOH, LiOH	NH_4OH, $Ca(OH)_2$, $Mg(OH)_2$
Salts	KI, NaCl, $CuSO_4$	PbI, $KHCO_3$, AgCl

The substances which are unable to conduct electricity in molten state or in aqueous solution form are called non electrolytes.

Example: Benzene, Glucose, Sucrose and Urea etc are non- electrolytes.

Test Yourself

- Define electrolyte?
- What are strong electrolytes?
- What are non-electrolytes?
- Identify strong and weak electrolytes from the following
  1. $HCl_{(aq)}$, 2. $KI_{(aq)}$, 3. $NaOH_{(aq)}$,4. $H_2S_{(aq)}$, 5. $CH_3COOH_{(aq)}$, 6. $NH_4OH_{(aq)}$

7.2.2 Electrolytic cells and Electrolysis:-

The electrolysis involves redox reactions and carried out in electrolytic cell. In electrolysis current passes through an electrolyte, due to this migration of positive and negative ions towards cathode and anode takes place. As a result of redox reaction ions are discharged at their respective electrodes.

The type of cell which uses electricity for a non spontaneous reaction to occur is called electrolytic cell.

An electrolytic cell consists of electrolyte in a vessel, electrodes and a battery.

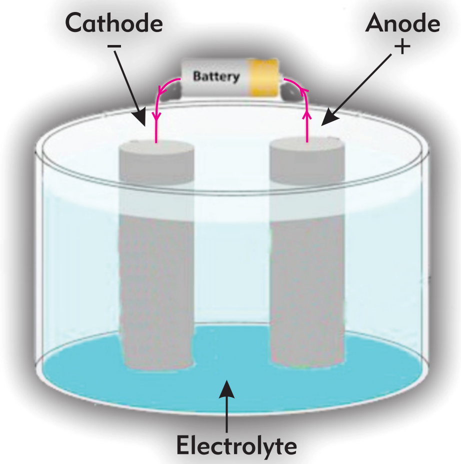


Fig 7.1 Electrolysis in electrolytic cell

The diagram of an electrolytic cell is shown in figure 7.1.

The figure shows that electrons from battery enter through cathode at which positive ions are reduced by accepting electrons. At anode negative ions loses electrons and undergoes oxidation. It means at cathode reduction occurs and oxidation takes place at anode.

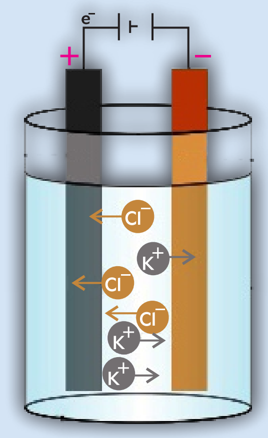
At Cathode $M^{+} + e^{-} \longrightarrow M$ (Reduction = Gain of electron)

At Anode $X^{-} \longrightarrow X + e^{-}$ (Oxidation = loss of electron)

Test Yourself

i. Sketch electrolytic cell for electrolysis of molten potassium chloride

ii. Identify cathode and anode, oxidation, reduction reaction, movement of electron from the following sketch of electrolytic cell.



Applications of Electrolytic cells:-

Important uses of electrolytic cell are given below.

I. It is used to prepare sodium metal from molten sodium chloride using the down's cell.

ii. It is used to prepare caustic soda (NaOH) from aqueous sodium chloride solution by Nelson's cell. It is also used to obtain chlorine gas.

iii. It is used to extract aluminum metal.

iv. It is used in electro refining of copper.

v. Electrolytic cells are used for electro plating of metals.

7.2.3 Faraday's law of Electrolysis:-

Michael Faraday was a British chemist who greatly contributed in the field of electrochemistry. He observed the quantitative relationship between current and the amount of substance collected at electrodes.

He conducted several experiments regarding electrolysis and put forward two laws of electrolysis based on his observation.

Faraday's First law of Electrolysis:-

It states that amount of any substance that is deposited or liberated at an electrode during electrolysis is directly proportional to the quantity of electricity passed through the electrolyte.

$$W \propto A \times t \qquad W \propto Q \qquad I = \frac{Q}{t} \qquad Q = It$$

or $$W = ZAt$$

In this equation W = weight of the substance deposited or liberated at electrode.

A = Current in Ampere

t = Time in second

Z = Electrochemical Equivalent

$$\text{Ampere (A)} \times \text{time}(t) = \text{Coulomb (C)}$$

If A = 1 Amp, t = 1 Sec then W = Z

Electrochemical equivalent is the weight of the substance collected at the electrodes when one coulomb of electric charge is passed through the electrolyte for 1 second.

Faraday's Second law of Electrolysis:-

The amount of different substances deposited or liberated due to passage of same quantity of current through different electrolytes are proportional to their chemical equivalent masses.

For an element

$$\text{Equivalent mass} = \frac{\text{Atomic weight}}{\text{valency}}$$

Example: Chemical equivalent of Al $= \frac{27}{3} = 9g$

Chemical equivalent of Ag $= \frac{108}{1} = 108$ g

Quantity of charge which deposits or liberates 1 gm equivalent weight of substance is called 1 Faraday (F).

1 F = 96500 Coulombs

Example:

Take three electrolytic solution of silver nitrate, copper sulphate and aluminum nitrate in three electrolytic cells and same quantity of current (96,500 coulombs) is passed through them. As a result 108 gm of silver, 31.75 gm of copper and 9 gm of aluminum are collected at their respective electrodes.

Batteries:-

We use lot of electrical devices having batteries as a source of electricity. A battery consists of group of galvanic cells connected in a series.

Examples of batteries include dry cell, lead storage battery, mercury battery etc.

Batteries are classified as primary (non rechargeable) and secondary (rechargeable) batteries.

Scientist are working for enhancing high energy, safety, recycle of batteries for mobile phones, transportation, computer technology etc.

Dry Cell:-

It is also known as Leclanche cell.

It is a type of primary cell which produce electricity using redox reaction between their chemical substances placed in it. It uses zinc as anode, magnese dioxide as cathode and aqueous ammonium chloride (NH_4Cl) or zinc chloride ($ZnCl_2$) as electrolyte. The cell diagram is given in Fig 7.2.

A copper cap is fixed on the top of the carbon rod for conduction of electricity.

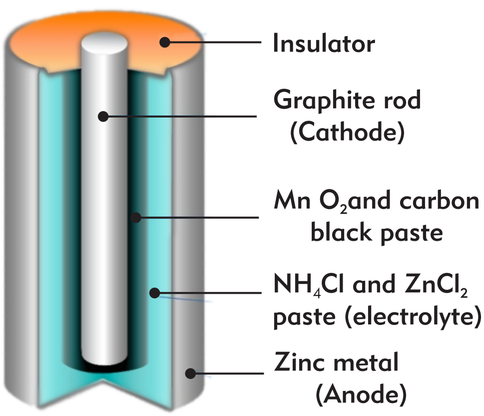


Fig 7.2 Dry cell

Zinc and graphite are then connected by a metal wire as a result following chemical reactions take place

Reaction at Anode $Zn \longrightarrow Zn^{+2} + 2\bar{e}$

Reaction at Cathode $2NH_4^+ + 2MnO_2 + 2e^- \longrightarrow Mn_2O_2 + 2NH_3 + H_2O$

It produces a potential of 1.5 volt.

Lead Storage Battery:-

A battery is a device which produces electricity through electro chemical reactions. Lead storage battery is an example of secondary cell in which chemical changes can be reversed. It has several voltaic cells connected in series .It contain lead plates which serve as anode and lead oxide (PbO_2) which acts as cathode. These electrodes are immersed in electrolytic solution of dilute sulphuric acid (H_2SO_4). Chemical changes during charging and discharging processes can be shown as

$$Pb + PbO_2 + 2H_2SO_4 \underset{\text{Charge}}{\overset{\text{Discharge}}{\rightleftharpoons}} 2PbSO_4 + 2H_2O$$

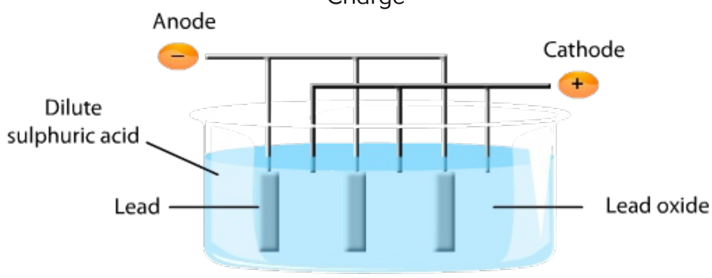


Fig: 7.3 Lead storage battery

Alloy Formation:-

Alloy is the mixture of metal with metal or metal with non metal. There are about 7000 alloys which are used for different purposes in the world.

Example: Brass is an alloy of Copper (Cu) and Zinc (Zn). Steel is a alloy of iron and carbon.

Alloy can be prepared by mixing elements in different proportions. In alloy it become difficult for layers of metal atoms to slide over each other. So alloy is harder and stronger than pure metal.

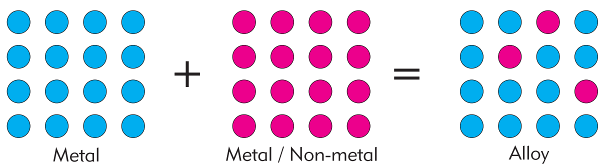


Fig 7.4 Alloy formation

Some important alloys are summarized below

Table 7.4

Name of Alloy	Components	Applications
Bell metal	Cu - Sn	Casting of bell
Brass	Cu - Zn	Door nobs and hard rails due to antibacterial nature, Hose nozzles, Stamping dies.
Bronze	Cu - Zn - Sn	Coins, medals, tools, etc.
Monel	Ni - Cu - Fe	Corrosion resistant containers
Duralumin	Al - Cu - Mg - Ni	Boat, Air craft etc
Solder	Sn - Pb - Cu - Sb	Joining electrical components into circuits.
Alnico	Fe - Al - Ni - Co	Magnets used in loud speakers
Amalgam	Hg - Ag - Cu - Zn	Dental filling
Cupronickel	Cu - Ni - Mn	Coins
Pewter	Sn - Cu - Pb - Sb - Bi	Ornaments
Sterling silver	Ag - Cu	Cutlery set, Medical tools
White gold (18 Karat)	Au - Pb - Ag - Cu	Jewelry

Do You Know?

24 karat gold is called 100% pure gold.

Addition of metals to gold form different colour.

Alloys of gold:

Yellow gold (22K) Alloy contains 91.67% gold with Ag, Cu, Zinc as other component.

Red gold (18K) contains 75% gold with Cu as other component.

White gold (18K) contains 75% gold with Cu, Ag as other component.

7.3 Corrosion and its Prevention:

Metals react with oxygen in presence of moisture and can form harmful metal oxide. These metal oxide layers are porous and expose metal for further reaction with oxygen to form harmful metal oxide. It is called Corrosion of metal.

7.3.1 Rusting of Iron

Corrosion of iron is an electro chemical process. Iron under goes redox reaction in presence of air or water to form iron (III) oxide ($Fe_2O_3 . nH_2O$) called rusting of iron. Rusted surface of iron provide no protection to underlying iron and eventually convert whole iron into reddish brown rust. Rusting occurs at different places of metal surface. A metal surface area of less moisture act as anode and oxidizes iron in this region .

$$Fe_{(s)} \longrightarrow Fe^{+2}_{(aq)} + 2\bar{e}$$

Metal surface with high moisture contents act as cathode and reduces atmospheric oxygen to OH^- ions.

$$O_2 + 2H_2O + 4\bar{e} \longrightarrow 4OH^-$$

The Fe^{+2} ions further reacts with oxygen to form rust iron (III) oxide($Fe_2O_3 . nH_2O$)

Prevention from corrosion:

All metals can be prevented from corrosion by following methods.

1- Alloying:

Formation of alloy prevents metal from corrosion by reducing its ability of oxidation.

Example: Iron (Fe) can be changed into stainless steel by mixing with chromium (Cr) and Nickle (Ni). Thus iron (Fe) is prevented from corrosion.

2- Metallic Coating (Electroplating)

All metals can be protected from corrosion by coating its surface with other metal like tin (Sn) or zinc (Zn). The coating of metal at the surface of other metal by electrolytic process is called electroplating. Metals like iron can be electroplated with chromium (Cr), Nickle (Ni) and silver (Ag).

3- Cathodic Protection:

It is applied to protect underground pipes tanks, oil rigs etc from corrosion by making these materials as cathode. The active metal like magnesium (Mg) or aluminum (Al) is used as Anode and connected with iron (Fe). These active metals itself oxidizes and prevent other metal from corrosion.

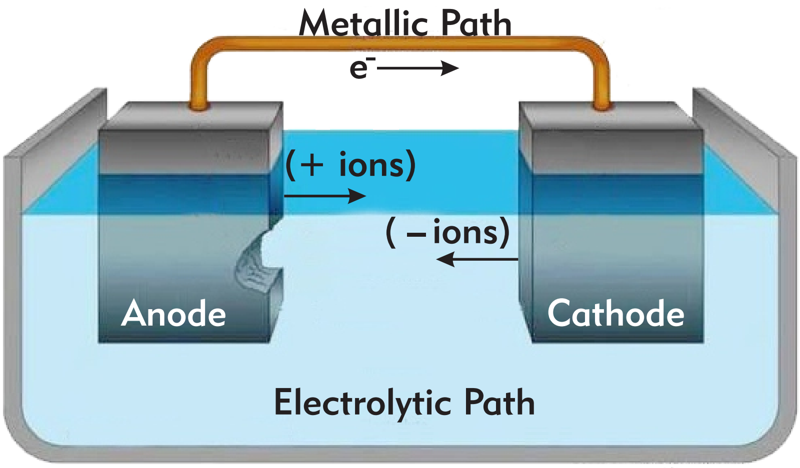


Fig: 7.5

4- Coating with paint:

A metal is commonly coated with paint to protect it from corrosion. Paint prevents the reaction of metal with oxygen moisture and other harmful chemical agents.

Test Yourself

- What is corrosion of metal?
- Name the methods which are used to protect metal from corrosion.
- How cathodic protection prevent metal from corrosion?

7.3.2 Electroplating on Steel:

The process of deposition of metal at the surface of other metal through electrolysis is called electroplating.

Tin Plating:

Steel spoon can be tin plated by using acidified tin sulphate as electrolyte. Tin (Sn) metal is used as anode and steel spoon is used as cathode. When current passes through electrolyte tin ions (Sn^{+2}) deposits at cathode as tin (Sn) metal. Tin (Sn) electrode is then changes into tin ion (Sn^{2+}).

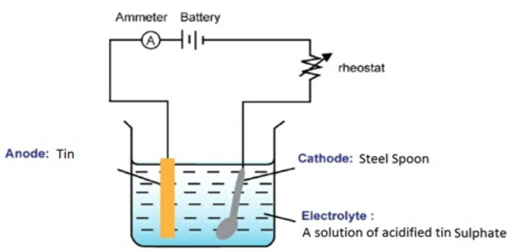


Fig 7.6 Tin plating of steel spoon

$$SnSO_{4(aq)} \rightleftharpoons Sn^{++} + SO_4^{2-}$$

$$Sn \longrightarrow Sn^{++} + 2e^- \quad \text{(Anode reaction)}$$

$$Sn^{2+} + 2e^- \longrightarrow Sn \quad \text{(Cathode reaction)}$$

Zinc Plating:

The process in which zinc is electrolytically coated at the surface of other base metal is called galvanizing. Potassium zinc cyanide is used as electrolyte to produce zinc ions(Zn^{+2}). Zinc (Zn) metal serves as anode and steel object is used as cathode. During electrolysis Zn^{++} deposits at cathode and Zinc (Zn) anode is then changes into zinc ion Zn^{2+}.

Following reactions occur during zinc electroplating.

$Zn \longrightarrow Zn^{2+} + 2\bar{e}$ (Reaction at anode)

$Zn^{+2} + 2e^- \longrightarrow Zn$ (Reaction at cathode)

Electroplating of Silver:

In this process silver (Ag) is coated electrolytically at the surface of steel or other metal. It is called silver plating. In this process aqueous solution of silver chloride (AgCl) is used as electrolyte to produce silver (Ag^+) ions. Silver (Ag) metal is used as anode and steel object like spoon used as cathode. Silver (Ag^+) ions are reduced at cathode by accepting electron. Silver anode loses electron and oxidized to silver (Ag^+) ion.

Following chemical changes takes place.

$$AgCl \rightleftharpoons Ag^+ + Cl^-$$

At cathode

$$Ag^+ + e^- \longrightarrow Ag$$

At anode

$$Ag \longrightarrow Ag^+ + e^-$$

Chromium Plating:

The process in which chromium (Cr) is coated electrolytically at the surface of other base metal is called chromium plating. Acidified chromium sulphate $Cr_2(SO_4)_3$ is taken as electrolyte. Chromium metal serves as anode and other metal object is used as cathode.

Following chemical changes take place in chromium plating.

$$Cr_2(SO_4)_{3(aq)} \rightleftharpoons 2Cr^{3+}_{(aq)} + 3SO_4^{2-}{}_{(aq)}$$

Reaction at anode:

$$Cr \rightleftharpoons Cr^{3+} + 3\bar{e}$$

Reaction at Cathode:

$$Cr^{3+} + 3\bar{e} \rightleftharpoons Cr$$

Chromium plated objects are used in Auto parts industries.

Society, Technology and Science:

Iron is reactive metal. It can react with food items and can spoil food.

Tin is non toxic, less reactive and resistant to corrosion. Tin can not react with organic acids or salt present in food so tin plated cans are used for beverages and for storing foods.

Silver is lustrous white metal. Many metal objects are silver plated to enhance its beauty and strength against corrosion.

Thin layer of silver on metal surface form durable layers. Thick coating of silver on metal surface is soft and gradually turns black, due to formation of silver sulphide(Ag_2S).

Concept Map

Electro Chemistry Deals with

- **Electro Chemical Cells**
  - **Electrolytic cell**
    - . it convert electrical energy into chemical energy
    - For Example
    - . Down cell
    - . Castner-Kellner cell
  - **voltaic (Galvanic) cell**
    - . it convert chemical energy into electrical energy
    - For Example
    - . Daniell cell
    - . Batteries
    - Electro Chemical cell contain electrodes
      - Cathode
        - . Reduction reaction occurs at cathode
        - . Positive ions move towards cathode
      - Anode
        - . Oxidation reaction occurs at Anode
        - . Negative ions move towards Anode
- **Electro chemical Reactions**
  - . These reactions convert electrical energy into chemical energy or vise versa.
  - . it involve redox reactions
- **Electrolytes**
  - . It consists of ions and conduct electricity
  - For Example
  - Aqueous solutions of Acids, bases or salts are electrolytes.

Oxidation Reaction

- . In this reaction oxidation number of an element increases by loosing electron.
- . introduction of oxygen or removal of hydrogen is also known as oxidation reaction.

Reduction Reaction

- . In this reaction oxidation number of an element decreases by accepting electron.
- . introduction of Hydrogen or removal of oxygen is also known as reduction reaction.

Summary

- Oxidation is the loss of electron by chemical substance.
- Reduction is the gain of electron by chemical substance.
- An electrolyte consists of free moving ions and conduct electricity.
- An electrode is an electrical conductor.
- The electrode at which oxidation occurs is called anode.
- The electrode at which reduction occurs is called cathode.
- Electrolysis is a process of migration of ions towards cathode and anode when current passes through en electrolyte.
- Oxidizing agent helps in oxidation by accepting electrons.
- Reducing agent helps in reduction by losing electrons.
- A galvanic cell converts chemical energy into electrical energy.
- An electrolytic cell uses electrical energy to drive a non spontaneous reaction.
- Corrosion of iron is called rusting.
- Corrosion of metal can be prevented by alloying, paint or electroplating metal by zinc, tin, silver, chromium etc.
- An alloy is a mixture of metal with metal or metal with non metal.

Exercise

SECTION- A: MULTIPLE CHOICE QUESTIONS

Tick Mark (✓) the correct answer

1. Alloy of Cu - Sn is called _______
   (a) Brass (b) Bronze
   (c) Monel (d) Bell metal
2. Which one is Alloy.
   (a) Graphite (b) Mercury
   (c) Steel (d) Water
3. Value of 1 Farady is equal to ____________.
   (a) 9.65 C (b) 9650 C
   (c) 96500 C (d) 965 C
4. Which one is non electrolyte.
   (a) Aqueous HCl (b) Aqueous NaCl
   (c) Molten KCl (d) Urea
5. Which one is oxidizing agent.
   (a) Al (b) H_2S
   (c) Cl_2 (d) NaH
6. Which one is reducing agent.
   (a) H_2SO_4 (b) HNO_3
   (c) Al (d) I_2
7. Which one forms weak electrolyte solution with water
   (a) HCl (b) KOH
   (c) NaCl (d) CH_3COOH
8. In dry cell _______ is used as cathode.
   (a) Zn (b) Cu
   (c) Sn (d) graphite
9. 1g equivalent weight of Al is equal to __________.
   (a) 9 g (b) 27g
   (c) 54 g (d) 1 g

10. Which one is correct statement.
    (a) Oxidation occurs at cathode.
    (b) Reduction occurs at anode
    (c) Reduction occurs at cathode
    (d) Ions lose electrons at cathode.

SECTION- B: SHORT QUESTIONS:

1. Define oxidation, reduction reactions with examples.
2. Why ionic compounds conduct electricity in molten or in aqueous solutions only?
3. What is electrolytic cell? Explain with diagram.
4. Define oxidizing and reducing agent with examples.
5. Examine the following chemical equations and identify.
   (i) Oxidizing agent (ii) Reducing agent
   (iii) Substance undergoes oxidation
   (iv) Substance undergoes reduction
   1. $Zn + Cl_2 \longrightarrow ZnCl_2$
   2. $Br_2 + H_2S \longrightarrow 2HBr + S$
   3. $2Ca + O_2 \longrightarrow 2CaO$
   4. $2Li + S \longrightarrow Li_2S$

Identify the Alloy.

Components	Alloy
Cu – Zn	
Cu – Al – Mg – Ni	
Cu – Zn – Sn	

SECTION- C: DETAILED QUESTIONS:

1. Describe the dry cell with diagram.
2. What is battery? How lead storage battery works?
3. Explain process of electrolysis in electrolytic cell.
4. What is Alloy? Explain its classification with examples.
5. What is rusting? How it can be prevented.
6. What is electroplating? How steel object can be electroplated with Tin, Zinc and Silver?
7. State and explain Faraday first and second law of electrolysis.

Chapter 8

CHEMICAL REACTIVITY

Time Allocation

Teaching periods	= 12
Assessment period	= 2
Weightage	= 12

Major Concepts

8.1 Metals

8.2 Non-Metals

STUDENTS LEARNING OUT COMES (SLO'S)

Students will be able to:

- Classifying the type of elements into metals, non-metals and metalloids.
- Draw flow chart diagram of classification of metals, non-metals and metalloids.
- Show how cations and anions are related to the terms metals and non-metals.
- Identify elements as an alkali metal or an alkaline earth metal.
- Analyze why alkali metals are not found in the free state in nature.
- Explain the differences in ionization energies of alkali and alkaline earth metals. .
- Describe the position of sodium in Periodic Table, its simple properties and uses. .
- Describe the position of calcium and magnesium in Periodic Table, their simple properties and uses.
- Differentiate between soft and hard metals (Iron and Sodium).
- Describe the inertness of noble metal.
- Identify the commercial value silver, Gold and Platinum.
- Compile some important reactions of halogens.
- Name some elements, which are found in uncombined state in nature.

Introduction

Medicines, plastics, glass, detergents etc are the products of chemical reactivity.

The property of substance to undergo chemical reaction with any material is called chemical reactivity.

Reactivity of metals depends upon its tendency to lose electron and that of non metals depends upon its tendency to accept electron.

Matter which undergoes chemical reactivity may be element, compound or mixture.

An element is always composed of similar atoms. Elements are further divided into metals, non-metals and metalloids.

Table 8.1

Metals	Non-Metals	Metalloids
❖ Tend to lose electrons in reactions ❖ Good conductors of heat and electricity ❖ Ductile ❖ Lustrous ❖ Strong ❖ Malleable ❖ Sonorous ❖ Oxides are basic in nature e.g. (Li_2O,Na_2O,KO_2,MgO).	❖ Tend to gain electrons in reactions with metals. ❖ Poor-conductors of heat and electricity ❖ Not ductile ❖ Often have dull appearance ❖ Oxides are acidic in nature (CO_2, SO_3, NO_2).	❖ Intermediate properties of metals and non-metals ❖ Boron (B), Silicon (Si), Germanium (Ge), Arsenic (As), Antimony (Sb), Tellurium (Te), Polonium (Po) and Astatine (At) are metalloids. ❖ Oxides may be acidic (B_2O_3, SiO_2) or amphoteric (AS_2O_3).

Do you know?

Most abundant elements found in air are

(1) nitrogen (2) Oxygen (3) Argon

Most abundant elements found in earth's crust are

(1) Oxygen (2) Silicon (3) Aluminum

Most abundant elements found in Universe are

(1) Hydrogen (2) Helium (3) Oxygen

Most abundant elements found in Human Body are

(1) Oxygen (2) Carbon (3) Hydrogen

8.1 METAL

The element which readily loses electron and easily form cation is termed as metal. A metal structure consists of metal ions joined by metallic bonds. All B group elements are metals and known as Transition Metals.

Some elements of A group are also metals.

Table 8.2

IA	IIA	IIIA	IVA	VA
Li	Be			
Na	Mg	Al		
K	Ca	Ga		
Rb	Sr	In	Sn	
Cs	Ba	Tl	Pb	Bi
Fr	Ra			

Elements of group IA are called Alkali metals.
Elements of group IIA are called Alkaline earth metals.
Metals of A - Group family are shown in table 8.2

Do you know?

- Beryllium (Be) is a light strong and highly toxic metal. Its small grain of 0.25 mg may kill a rat.
- Most abundant metal is aluminum (Al).
- Most useable metal is iron (Fe).
- Most reactive metal is cesium (Cs).
- The lightest metal is lithium (Li).
- The heaviest metal is osmium (Os).
- Most malleable, ductile metals are gold(Au) and silver (Ag).

8.1.1 Electropositive Character (Cation formation)

Metals are highly electro positive, due to this property they easily lose their valance shell electrons. When an atom or a molecule loses electron then it changes into positively charged ion known as Cation.

$$Al \longrightarrow Al^{3+} + 3e^-$$

Electro positive character of metals increases down the group with increasing atomic size. Alkali metals have large atomic size and low ionization potential values. The nucleus force on valence shell is very weak so they can lose their valence electron easily. Hence they are highly reactive, highly electro positive, powerful reducing agents and cannot exist free in nature.

The valence shell electronic configuration of alkali metals is ns^1.

Alkali metals lose one electron and form monovalent cation.

Example Li^+, Na^+, Rb^+, Cs^+

The valence shell electronic configuration of alkaline earth metals is ns^2.

Do you know?

Alkali and Alkaline earth metals can be identified by the flame test.

Name	Symbol	Colour of Flame
Lithium	Li	Bright Crimson (Bright Red)
Sodium	Na	Golden Yellow
Potassium	K	Violet
Rubidium	Rb	Lilac (Dark Red)
Cesium	Cs	Bright Blue
Beryllium	Be	White
Magnesium	Mg	Bright White
Calcium	Ca	Brick Red
Strontium	Sr	Crimson Red
Barium	Ba	Green

Test Yourself

- Which metal is found in liquid state?
- Identify the alkaline earth metals from the following elements.

  Bi, Br, Ba, B, Se, Si, Sb, Sr
- Write few properties of metals.
- Write some properties of non-metals.

Ionization Energy of Alkali and Alkaline Earth Metal

The removal of electron from an element requires energy which is known as Ionization Energy

$$\text{Atom} + \text{Energy} \longrightarrow \text{Cation} + e^-$$

Table 8.3

Ionization Potential in KJ / mole

Reactivity Increases ↓

I-A	II-A
Li = 520	Be = 899
Na = 495	Mg = 738
K = 419	Ca = 520
Rb = 403	Sr = 549
Cs = 376	Ba = 309

Reactivity Increases ↓

Ionization energies values decreases with increasing atomic size and vice versa.

The Alkali metals and Alkaline earth metals show increasing trend of reactivity down the group because their atomic size increases down the group.

Since Alkali metals have low Ionization Energy values than Alkaline earth metals, Alkali metals are highly reactive than Alkaline earth metals.

Alkali Metals and Alkaline earth metals have low values of Ionization Energies due to which they easily lose their valence electron and form Cation. Thus they are highly reactive.

8.1.2 Comparison of reactivity of Alkali and Alkaline Earth Metals:

The general comparison of reactivities of IA and IIA group elements is shown bleow:

Alkali Metals (IA)	Alkaline Earth Metals (IIA)
They are highly reactive than (IIA) group elements due to low ionization energy.	They are less reactive than (IA) group elements due to high ionization energy.
They form monovalent cation(M^+)	They form divalent cation(M^{2+})
They immediately tarnish in air and form metal oxide. $K + O_2 \rightarrow K_2O$	They react with oxygen on heating. $2Mg + O_2 \rightarrow 2MgO$
They react violently with halogens $2Na + Cl_2 \rightarrow 2NaCl$	They react slowly with halogens $Ca + Cl_2 \rightarrow CaCl_2$
They react with water vigorously at room temperature and form strong alkaline solution $2K + 2H_2O \rightarrow 2KOH + H_2$	They react with water less vigorously and form alkaline solution $Mg + H_2O \rightarrow MgO + H_2$ $MgO + H_2O \rightarrow Mg(OH)_2$
Their oxides and hydroxides are more basic than those of IIA group elements.	Their oxides and hydroxides are less basic than those of IA group elememts.
They do not form metal carbides.	They form metal carbides on heating. $Ca + 2C \rightarrow CaC_2$

Position of Alkali metals and Alkaline earth metals in periodic table is also useful to explain their reactivity. Detailed discussion about reactivity of sodium, magnesium and calcium is given below:

Position, Properties & Uses of Some Metals

Sodium (Na)

Position:

It is sixth most abundant element and constitutes 2.87% of earth's crust. It belongs to IA group, 3rd period of periodic table.

Properties:

Sodium is silvery white alkali metal. It melts at 97.8°C and boils at 882.8°C. It is soft and can be cut with Knife due to weak metallic bonding between their atoms. It violently reacts with H_2O water and form Sodium Hydroxide and Hydrogen gas, so it is kept in Kerosene oil to prevent its reaction with moisture.

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

Sodium metal has shining surface but its appearance becomes dull due to action of air.

$$2Na + O_2 \rightarrow Na_2O$$

Sodium reacts with halogens to form sodium halide.

$$2Na + Br_2 \rightarrow 2NaBr$$

Sodium reacts with sulphuric acid to form H_2 gas

$$2Na + H_2SO_4 \rightarrow Na_2SO_4 + H_2$$

Uses:

It is an excellent heat transfer fluid so it is used as coolent in nuclear reactors.

It is used in Detergent preparation.

It is used in Street lights and gives yellow colour.

It is used as Reducing agent in the extraction of Calcium, Zirconium and Titanium.

Some common compounds of Sodium and their uses are mentioned below:

Table 8.4

Compound	Formula	Uses
Soda Ash	Na_2CO_3	Used as water softener
Baking Soda	$NaHCO_3$	Used in Baking Powder, Health Salt, Beverages
Table Salt	$NaCl$	Food Items
Sodium Nitrate	$NaNO_3$	Used as fertilizer and in Dynamite

Magnesium (Mg)

Position:

It is 8th most abundant element found in earth's crust. Magnesium belongs to II-A group and 3rd period of periodic table.

Properties:

It is a grey-white metal. Its name is derived from Magnesia, a district in Greece. It melts at 650°C and boils at 1090°C. Magnesium reacts with water and releases Hydrogen gas

$$Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$$

Magnesium fire cannot be extinguished with water because H_2 gas is highly flammable and intensifies the fire. Magnesium fire can be extinguished by using dry sand.

Uses:

It is used in flares and photographic flash bulbs because it burns to produce brilliant white light. Magnesium hydroxide are used as an Antacid. It is used for manufacturing of Mobile Phones, Laptop and Tablet Computers because of light weight and electrical properties.

The use of Magnesium reduces the weight of vehicle by replacing steel components of a vehicle.

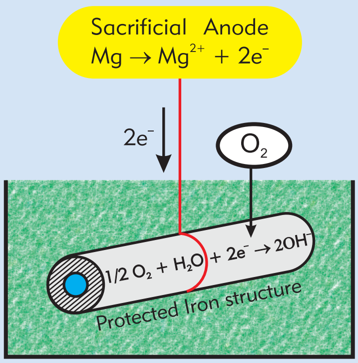
Magnesium alloys are used in aviation industry, space crafts and missile because they are light weight and remain stable at high temperature.

Magnesium can be changed into intricate (twisters, knotty) shapes, so it is used in tenis rackets and handles of archery bows.

Do you know?

Magnesium is used in **cathodic protection (CP)**

Magnesium is easily oxidized as compared to iron so magnesium is used as anode and underground metallic pipelines become cathode to protect from corrosion.



Following reaction take place

$$Mg \rightarrow Mg^{2+} + 2e^- \quad \text{(Oxidation) (At Anode)}$$

$$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O \quad \text{(Reduction) (At Cathode)}$$

Calcium (ca)

Position:

It is 5th abundant metal in earth's crust. It belongs to IIA group and 4th period.

Properties:

It is silvery white soft metal. It melts at 851°C and boils at 1484°C.

Uses:

Calcium is essential for healthy teeth and bones.

Calcium forms useful compounds which are mentioned below:

Table 8.5

Compound	Formula	Uses
Slaked lime	$Ca(OH)_2$	As soil conditioner, used in water treatment to reduce acidity. Used in steel industry to remove impurities from Iron ore
Gypsum	$2CaSO_4.H_2O$	It is used as component in construction of buildings. It is used medically in plaster for setting broken bones.
Calcium hypochlorite	$CaOCl_2$	It is used for sterilization of water in swimming pool.
Calcium tungstate	$CaWO_4$	It is used in Luminous paints.
Limestone	$CaCO_3$	As source of CO_2, In Cement industry.

Do you know?

"Calcium Light"

When a popular personality becomes the center of public attention then it is said to be in the Limelight. Previously musical halls, theaters stages, were lit by chemical called lime (CaO), using Oxy-Hydrogen flame. As a result bright light was produced which is known as Calcium light.

This light increases the visibility of audience to enjoy the performance of Actor on stage.

Test Yourself

- Write uses of baking soda, bleach and sodium nitrate.
- Enlist the uses of magnesium.
- Write uses of Slaked lime , Gypsum and calcium tungstate.

Soft and Hard Metals

Metals may be soft or hard. The hardness of metal is the resistance of metal to be scratched. It is measured in moh scale.

The metals which are scratched easily are called soft metals.

Alkali metals like; Sodium(Na), Potassium (K) and Rubidium (Rb) are soft metals.

Metals which show strong resistance towards scratching are called hard metals.

Nickle (Ni), Iron (Fe), Tungsten(W) are hard metals.

soft and hard metals can be differentiated with reference to sodium (Na) and iron (Fe) as follows.

Table 8.6

Sodium	Iron
It is a soft metal of group IA	It is a hard metal VIIIB
It has large atomic size	It has smaller ionic radii
It has low value (0.5) on moh scale	It has high value (4.5) on moh scale
It has weak metallic bonding so it is a soft metal	It has strong metallic bonding so it is hard metal.
It can be cut easily with knife.	It is hammered to form sheets and wires.
It is light due to its low density (0.971 g/cm^3).	It is heavier metal due to its high density (7.87 g/cm^3).
It has low melting and boiling point values (melting point = 98°C, boiling point =882.8°C)	It has high melting and boiling point values (melting point = 1535°C, boiling point = 2450°C)

Test Yourself

- Define Soft metals with examples.
- Define hard metals with examples.
- Write melting point, boiling point, density and moh values of sodium and iron.

Do you know?

Hardness of metals and other material is measured by Moh scale. It is devised by Frederich Mohs in 1812. It is based on scratch resistance of different metals and other materials. Hardness of plastic, Lead pencil is 1, and that of Diamond is 10 on Moh scale.

Values of Hardness of few metals on Moh scale

Li	Na	K	Rb	Cs	Ni	Fe	W
0.8	0.5	0.4	0.3	0.2	4	4.5	7.5

8.1.3 Inertness of Noble Metal

Noble metals include Gold (Au), Silver (Ag), Platinum (Pt), Iridium (Ir), Osmium (Os), Rhodium (Rh), Ruthenium (Ru), Palladium (Pd).

Nobel metals are less electro positive so they are difficult to oxidize. Therefore they show no reaction with atmospheric gases and resist corrosion. This helps noble metals to maintain their appearance so noble metals like Ag, Au, Pt are used to make ornaments.

Do you know?

Two pieces of pure un-coated metals permanently stuck together in space because there is no oxygen in space and hence no oxidation reaction occurs. The oxidized layer on metals serves as a barrier and prevent adherence of metals.

Commercial Value of Silver (Ag), Gold (Au), Platinum (Pt)

SILVER (Ag): - It is widely used in society. It is used in Jewelry, decorative items and Silver tableware because it does not tarnish and maintain its silvery shiny appearance. It is used to make mirror because it is best reflector of visible light. Silver forms compounds of significant importance.

Silver Nitrate ($AgNO_3$) or Lunar caustic is used in detection of Halogen. Light sensitive material AgBr and AgI are used in Photographic films.

GOLD (Au): - Gold has importance in our society. It is used in Jewellery because it has very high luster, yellow colour and tarnish resistance.

Gold is used in Electronic components because it is highly efficient conductor of current and cannot be corrode.

Gold is used in connecting wires, connection strips, switches and relay contacts to make electronic devices highly reliable. Therefore, Gold is used in cellphones, global positioning systems, Calculators etc. Gold is used in Laptop Computers for rapid and accurate transmission of digital information. It is used in dentistry because it is chemically inert, non-allergic and easy for dentist to work. Gold coated polyester films are used in space vehicles to reflect infrared radiation and stabilize the temperature of space vehicle. The helmet of Astronaut is coated with thin film of gold which reflect intense Solar radiation and protect eyes, skin of astronaut. Glass surface coated with gold will reflect solar radiations outward and keep the Buildings cool in summer. It also reflects internal heat inward and keeps the Building warm in winter.

Gold symbolizes purity, beauty and stability so it is used in making medals, trophies awards etc.

PLATINUM (Pt): - It is a silvery white corrosion resistance metal. It is paramagnetic transition metal.

It is used in chemical reactions as catalyst.

Reaction: -

It is used as catalytic converter in vehicles. It helps the complete combustion of Hydrocarbons and reduces the emission of air pollutants.Price of precious metal is fixed according to its weight. Its density is more than Gold. So it is more expensive than Gold.

Test Yourself

- Write names and symbols of few noble metals?
- Why helmets of astronauts are coated with thin film of gold?
- Why glass surface are coated with gold?
- Why gold is used in jewelry?
- Why platinum is used as catalytic converter in vehicles?

Do you know?

Liquid Metal

Mercury is the only metal found in liquid state. It belongs to sixth period and II B group of modern periodic table.

It has lowest melting point among all metals

It forms alloys with other metals which are known as amalgam.

For example, tin amalgam is an alloy of tin and mercury.

Its alloy with silver and tin is used as dental filling.

Mercury is used in thermometer and barometer.

Gaseous mercury is used in street light and fluorescence lamps.

8.2 NON - METALS:

Non-metals are the elements which have greater tendency to accept electron. Non-metals are placed at upper right portion of periodic table as shown in table 8.7.

Table 8.7
Non-Metal of Periodic Table

IA	IVA	VA	VIA	VIIA	VIIIA
1 H Hydrogen 1.00797					2 He Helium 4.0026
	6 C Carbon 12.01115	7 N Nitrogen 14.067	8 O Oxygen 15.9094	9 F Fluorine 18.9094	10 Ne Neon 20.180
		15 P Phosphorus 30.9738	16 S Sulphur 32.064	17 Cl Chlorine 35.453	18 Ar Argon 39.948
			34 Se Selenium 78.98	35 Br Bromine 79.904	36 Kr Kryptos 53.80
				53 I Iodine 126.9044	54 Xe Xenon 131.30
					86 Rn Radon (222)

Non-metals are non malleable; non ductile, dull in appearance, non-sonorous, poor or nonconductors of heat and electricity.

Most of the non-metals are gases

For example: H, N, O, F, Cl and VIIIA group non-metals are gases.

Bromine (Br) is the only non-metal found in liquid state.

Some non metals like S, P, Se, I are solids.

Test Yourself

- Describe the properties of non-metals.
- Identify the VIIIA group elements from the following

  N, Na, Ni, Ne, Ar, At, He.
- Write names and symbols of Non metals of VA group elements.
- Which group contain non metals in gaseous state only?

Do you know?

Fluorine gas is a yellow color non-metal.

Chlorine gas is a green non-metal.

Iodine is a lustrous purple color non-metal.

Diamond is hardest non-metal.

8.2.1 Electronegative Characteristics:

A Non-metal has property to accept electron easily and form Anion. It is called electronegative character.

$$x + e^- \rightarrow x^-$$

Negative ions move towards anode during electrolysis so they are termed as anions. The number of negative charges on anion shows the excess number of electrons as compared to number of protons.

Electronegative character increases across the period because Atomic size decreases and nuclear charge density increases. It decreases down the group due to increasing Atomic size.

Halogens accept electrons easily due to their high electronegative character.

$$Cl + e^- \rightarrow Cl^-$$

Non-metals form acidic oxides which react with water vapors of atmosphere and cause acid rain.

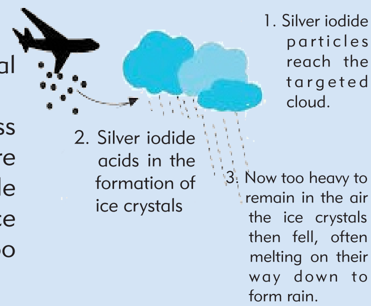
$$SO_2 + H_2O \rightarrow H_2SO_3 \text{ Sulphurous Acid}$$
$$SO_3 + H_2O \rightarrow H_2SO_4 \text{ Sulphuric Acid}$$
$$CO_2 + H_2O \rightarrow H_2CO_3 \text{ Carbonic Acid}$$

Do you know?

Artificial Rain

Dr. Vincent J. Schaefer in 1946 successfully created artificial clouds in a chilled chamber.

Artificial Rain can occur through cloud seeding. In this process chemicals like Silver Iodide (AgI) or Dry ice (Solid CO_2) are spread over clouds. As a result a super cooled water molecule condenses rapidly around these chemicals and forms ice crystals. When these ice crystals grow big and become too heavy then they fell downward and change into rain.



8.2.2 Comparison of Reactivity of the Halogen

Halogen belongs to VII-A group and consists of Fluorine (F), Chlorine (Cl), Bromine (Br), Iodine (I) and Astatine (At). Halogens exist in Molecular form. The reactivity of halogens decreases down the group because atomic size increases and electro negativity decreases down the group.

1. Halogens act as oxidizing agent, because they easily accept electron.

$$2Na + Cl_2 \rightarrow 2NaCl$$
$$2K + I_2 \rightarrow 2KI$$

Power of Halogens as oxidizing agent decreases in the following order

$$F_2 > Cl_2 > Br_2 > I_2$$

It means Fluorine can displace other Halogens due to its highest oxidizing power.

2. More reactive Halogen can displace less reactive Halogen form a solution of its salt.

Example

$$2KI + Cl_2 \rightarrow 2KCl + I_2$$

Chlorine is more reactive Halogen than Iodine so Chlorine displaces Iodine. The mixture turns reddish brown due to liberation of Iodine.

3. Fluorine is more reactive, therefore it can replace other halogens.

$$2KCl + F_2 \rightarrow 2KF + Cl_2$$
$$2KBr + F_2 \rightarrow 2KF + Br_2$$

4. Reaction of Hydrogen with Halogen form Halogen acid.

$$H_2 + X_2 \rightarrow 2HX \quad \begin{matrix} (X_2 = F_2, Cl_2, Br_2, I_2) \\ (HX = HF, HCl, HBr, HI) \end{matrix}$$

The Acidic strength of Halogen Acid decreases in the following order

$$HI > HBr > HCl > HF$$

HI is very strong acid because HI easily breaks up and forms H^+ Ion in water due to weak Covalent Bonding.

HF is very weak acid because it has strong Covalent Bonding so it does not break up easily to form H^+ Ion in water. H^+ Ions reacts with water molecule to forms Hydronium (H_3O^+) Ion.

Do you know?

Enamel is the outer layer of our teeth it is mainly composed of calcium carbonate ($CaCO_3$) and hydroxy apatite $[Ca_3(PO_4)_2]$. $Ca(OH)_2$. Fluoride (F^-) ions of toothpaste replaces hydroxide (OH^-) ions of hydroxy apatite and form flouro apatite. This replacement makes it more resistant against tooth decay.

Skills:

Qualitative analysis of Cations like Zn^{+2}, Mg^{+2}, NH_4^{+1} , Ca^{+2} and Ba^{+2} can be identified by following test.

EXPERIMENT

Table 8.8

Test for Zn^{2+}

Experiment	Observation	Result
Salt Solution + NH_4OH solution.	White ppt	Zn^{2+} may be present
White ppt + excess NH_4OH or NaOH solution.	White ppt is dissolved and form clear solution	Zn^{2+} ion is present

Test for Mg^{2+}

Experiment	Observation	Result
Salt Solution + NaOH or NH_4OH solution.	White ppt	Mg^{2+} ion may be
White ppt + NaOH/ NH_4OH solution.	White ppt insoluble in excess of NaOH or NH_4OH	Mg^{2+} is present

Test for NH_4^+

Experiment	Observation	Result
Portion of Aqueous solution of salt + NaOH solution warm	Pungent of NH_3 Gas release	NH_4^+ Ion is present

Test for Ca^{2+} & Ba^{2+}

Experiment	Observation	Result
Heat Nichrome wire until flame is no longer coloured.	Apple Green colour flame	Ba^{2+} is present
Dip the loop of wire into water and then in unknown salt. Heat the wire on flame.	Brick Red colour flame	Ca^{2+} is present

Qualitative analysis of Anions like CO_3^{2-}, Cl^-, I^-, SO_4^{2-} and NO_2^- can be identified by following test.

Test for CO_3^{2-}

Experiment	Observation	Result
Sample of Solid + Dilute Mineral Acid	Bubbles comes out which turns lime water milky	CO_3^- is present

Test for Cl^-

Experiment	Observation	Result
Few ml of salt solution + dilute HNO_3 + $AgNO_3$	White ppt	Cl^- may be present
White ppt + NH_4OH solution	White ppt soluble in NH_4OH	Cl^- Ion is present

Test for I^-

Experiment	Observation	Result
Few ml of salt solution + few drops of dilute HNO_3 + few drops of $AgNO_3$	Yellow ppt	I^- ion may be present
Yellow ppt + NH_4OH solution	Yellow ppt insoluble in excess NH_4OH	I^- ion is present

Test for SO_4^{2-}

Experiment	Observation	Result
Few ml of Sample + dilute HCl + few drops of $BaCl_2$ or Few drops of Sample of dilute HCl + few drops of lead nitrate $Pb(NO_3)_2$ soultion	White ppt	SO_4^{2-} Ion may be present
White ppt + dilute HCl	White ppt insoluble in excess HCl	SO_4^{2-} Ion is present

Test for NO_2^-

Experiment	Observation	Result
Small amount of salt + few drops of dilute H_2SO_4 solution	Reddish Brown vapours evolve	NO_2^- Ion is present

Concept Map

Chemical Reactivity
(Chemical interaction between chemical substances)

↓

Involve matter which may be

↓

Non-metal

- Easily accept electron and form negative ion (electro negative character).
- Halogens are non-metals.
- Halogens act as oxidizing agents.
- Elements of group VIIIA are non-metals.

Metal

- Easily loses electron and form cation (electropositive character).
- IA, IIA group elements are highly electro positive form basic oxides and act as reducing agents.
- Nobel metals are less electropositive so they are corrosion resistant and difficult to oxidize.
- Metals resistance against scratching is called hardness

↓

- Hardness is measured in moh scale.
- Alkali metals are soft metals.
- Metals like iron (Fe), nickle (Ni), tungsten (W) are hard metals.

Metalloids

- Elements having intermediate character of metals and non-metals.
- Can form acidic or amphoteric oxides.

Summary

- Metals have greater tendency to lose electron.
- Non-metals have greater tendency to accept electron.
- Oxides of metals are basic because they produce basic solution with water.
- Oxides of Non-metals are Acidic because they form Acidic solution with water.
- Ionization Energy decreases and Electro positivity increases down the group.
- Group I-A elements are called Alkali Metals.
- Group II-A elements are called Alkaline earth metals.
- Metals of Group I-A, II-A are powerful reducing agents.
- Noble metals like Platinum, Silver, Gold etc are difficult to oxidize.
- VII-A group elements are Non-metals and act as oxidizing agents.
- VII-A group elements are known as Halogens.
- Halogen reacts with metals and form salt.
- Elements of group VIIIA are non-metals and consist of gases.
- Metalloids are elements which have properties intermediate between metals and Non-Metal elements.

  For example B, Si, Ge, As, Sb, Te are metalloids.

Exercise Questions

SECTION- A: MULTIPLE CHOICE QUESTIONS

Tick Mark (✓) the correct answer

1. Which one metal belongs to Alkaline earth metal?
   (a) B (b) Bi (c) Br (d) Ba
2. Which one is Barium
   (a) Bi (b) Be (c) Ba (d) Br
3. Chlorine can be displaced by __________ .
   (a) F (b) Br (c) I (d) At
4. Which one is strong acid?
   (a) HF (b) HCl (c) HBr (d) HI
5. Which Halogen exists in liquid state?
   (a) F_2 (b Cl_2 (c) Br_2 (d) I_2
6. Non-metals of _________ group are gases
   (a) VI-A (b) VII-A (c) VIII-A (d) VIII-B
7. Which one is Metalloid?
   (a) Br (b) Se (c) S (d) Sr
8. Which one of the following act as oxidizing agent ?
   (a) Be (b) Mg (c) Na (d) Cl
9. Which gas can turn lime water milky?
   (a) O_2 (b) NO_2 (c) CO_2 (d) N_2
10. Which compound is known as lunar caustic?
    (a) KNO_3 (b) $AgNO_3$ (c) NaOH (d) $NaNO_3$

SECTION- B: SHORT QUESTIONS:

1. Identify the elements as Metals, Non-metals and Metalloids from the following elements:-

Elements	Metals	Non-Metals	Metalloids
C, Ca, Sb, S, Sr, Si, K, P, Ba, Ge			

2. What are Alkali metals? Write names and symbols of alkali metals?
3. What are Alkaline Earth Metals? Write names and symbols of Alkaline earth metals?
4. What are Halogens? Write names and symbols of Halogens?
5. Write names and symbols of any three metalloids.
6. Define following with examples: - 1. Cation 2. Anion
7. Explain why alkali metals do not exists free in nature?
8. Explain inertness of noble metals.
9. Write names and symbols of some noble metals
10. Explain electronegative character of non-metals.

SECTION- C: DETAILED QUESTIONS:

1. Explain importance of Silver.
2. Explain importance of Gold.
3. Explain the experiment to test Cl^- and I^- Ions.
4. Explain Electropositive character of metals.
5. Explain the position of Magnesium in periodic table and its importance.
6. Explain the position of Sodium in periodic table and its importance.
7. Arrange the following Halogen Acids in increasing order of their Acidic strength: HBr, HCl, HI, HF
8. Explain Electronegative character of non-metals.
9. Differentiate between sodium and iron as soft and hard metal.
10. Discuss the reactivity of Halogens.