

IX CHEMISTRY

Complete Notes

MUHAMMAD HASSAM

CHAPTER # 01

INTRODUCTION OF
CHEMISTRY

CHEMISTRY:

“It is the branch of science which deals with the composition and properties of matter and the chemical changes involved in it”.

BRANCHES OF CHEMISTRY

The field of chemistry has been divided into the following main branches.

1. Physical chemistry.
2. Organic chemistry.
3. Inorganic chemistry.
4. Analytical chemistry.
5. Bio – chemistry.
6. Industrial chemistry or applied chemistry.
7. Nuclear chemistry.
8. Environmental chemistry.
9. Polymeric chemistry.

PHYSICAL CHEMISTRY:

“It is a branch of chemistry in which deal with the forces and principals involved in the combination of atoms and molecules.”

ORGANIC CHEMISTRY:

“A branch of chemistry in which we study of carbon containing substances and the synthesis of such materials.”

INORGANIC CHEMISTRY:

“It is a branch of chemistry in which we study about the element and their derivatives compounds except those of carbon hydrogen compound and their derivatives.”

ANALYTICAL CHEMISTRY:

“A branch of chemistry in which we study of the methods and techniques used to determine the kind and quality of various components in a given substance.”

BIO-CHEMISTRY:

“A branch of chemistry that deals with the compounds of living organisms I e, plants, animals, and their metabolism in the living body.”

INDUSTRIAL CHEMISTRY:

“In this branch we study of different chemical process involved in the industry for the manufacture of synthetic products like glass, cement, medicine etc.”

NUCLEAR CHEMISTRY

“In this branch of chemistry that deals with the study of the changes occurring the nuclei of the atoms acquired by emission of radiation.”

ENVIRONMENTAL CHEMISTRY:

“It is the study of the interaction of chemical materials and their effects on environmental of animal and plants.”

POLYMERIC CHEMISTRY:

“It is the branch of chemistry that deals specially with the study of polymerization and the products obtained through the process of polymerization such as plastics, synthetic fiber, papers etc.”

HISTORY OF CHEMISTRY:

Over the centuries, chemistry has undergone remarkable progress in every field. All these developments were empirical that is why they were achieved by oral and error and not on the basis of any system atic study.

History of chemistry is divided in to three periods.

1. The Greek period.
2. The Muslims period.
3. The Modern period.

1. THE GREEK PERIOD:

“The Greek philosopher were the first the developeraliting to chemistry. They introduced the concept elements, atoms, shapes of atoms and chemical combination. They believed that all matter was dividing from four elements earth, fire, air and water. These elements had properties such as hot, dry, cold and wet.”

The Romans develops the chemical arts still further. They improved metallurgical process and introduced the enameling of pottery. Their works were all empirical.

2. THE MUSLIM PERIOD:

“The Muslim period was from 600 1605 AD in the history of chemistry and is known as the period of alchemists. The Muslim scientists contributed richly towards may branches of science, especially field of chemistry and introduced new methods and experimentation.”

Modern scientific knowledge based on the contribution of these Muslim scientists,

- | | |
|------------------------|-------------------|
| 1. JABIR – BIN – HAYAN | (721 – 803 A.D) |
| 2. AL – RAZI | (862 – 930 A. D) |
| 3. AL – BERUNI | (973 – 1048 A. D) |
| 4. IBNE – SINA | (980 – 1037 A. D) |
| 5. IBN – UL – HAITHAM | |

3. THE MODREN PERIOD

“Modern chemistry began in the 17th and 18th century and the result of experimentation. The beginning of 19th century is marked by the development of Dalton’s atomic theory. T he do the century is characterized by tremendous growth of chemical knowledge.”

In modern age, these scientists contributed in the field of chemistry.

- | | |
|--------------------|-------------------|
| 1. Robert boyle | (1627 – 1691 A.D) |
| 2. J. Precisely | (1733 – 1804 A.D) |
| 3. Mendel eve | (1824 – 1907 A.D) |
| 4. Arrhenius | (1859 – 1927 A.D) |
| 5. Henry Becquered | (1852 – 1908 A.D) |
| 6. J. J Thomson | (1856 – 1940 A.D) |
| 7. Neil bohr | (1885 – 1962 A.D) |
| 8. E. Rutherford | (1891 – 1937 A.D) |

THE SCIENTISTS APPROCH IN CHEMISTRY

The observed natural phenomenon had to be explained by different scientific methods in synthetic manner. The scientific approach in chemistry can be described through observation hypothesis, theory and law.

1. **OBSERVATION:**

“Observation is the basis that for celebrating a phenomenon but it may vary from person to person according to his own skill. It is one of the scientific approaches in chemistry.”

2. **HYPOTHESIS:**

“After observation of a phenomenon, a scientist ponders over it and carries out relevant experiments and arrives at a possible explanation of the nature, of the phenomenon. This explanation which is a trial idea is known as hypothesis.”

3. **THEORY:**

“When the hypothesis is supplied by a large number of different types of observations and experiments by other workers on the same phenomenon, then a theory is developed.”

4. **LAW:**

“When a theory is tasted again and again and is found fit according to the facts, giving valued prediction is called as scientific law.”

CHAPTER # 02

CHEMICAL COMBINATION

CHEMICAL COMBINATION:

“The combination in which two or more elements combine chemically and they lose their original properties called chemical combination”.

LAWS OF CHEMICAL COMBINATION:

There are several laws that the composition of matter and chemical reaction, known as laws of chemical combination. These laws are:

- (i) Law of Conservation of mass (or matter)
- (ii) Law of definite proportion (or constant composition)
- (iii) Law of multiple proportion
- (iv) Law of reciprocal proportion

(i) THE LAW OF CONSERVATION OF MASS

Introduction:

It is one of the most important scientific laws. This law was put forward by a French chemist, Lavoisier in 1785.

Statement:

“Matter is neither created nor destroyed during a chemical reaction but changes from one form to another”.

Explanation:

The law is now accepted as the basis of all chemical reactions. Its validity can be easily demonstrated in the laboratory by a precipitation reaction. It should be remembered that during a chemical reaction, neither the number of atoms nor the types of atoms change, so the mass is conserved.

Example: The law of conservation of mass is practically proved by the example of Lavoisier's experiment.

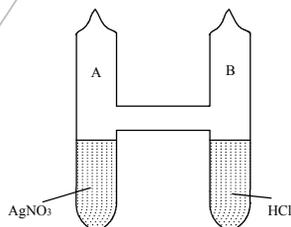
LANDOLT'S EXPERIMENT INTRODUCTION:

Introduction:

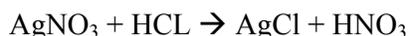
German Chemist H. Landolt was among the first chemists to demonstrate the validity of the statement of this law.

Detail of experiment:

In this experiment Landolt took an H-shaped tube and filled the two limbs A and B, with silver nitrate (AgNO_3) in limb A and hydrochloric acid (HCl) in limb B. The tube was sealed so that the material could not escape outside.



The tube was weighed initially in a vertical position so that the solutions should not intermix with each other. The reaction proceeds and reactants were mixed by inverting and shaking the tube. The tube was weighed after mixing on the formation of white precipitate of (AgCl). He observed that the weight remains the same.



The total mass of the substance before the reaction is equal to the total mass of the substance after the reaction.

(ii) THE LAW OF DEFINITE PROPORTIONS:

Introduction:

The second law of chemical combination was proposed by French chemist Louis Proust in 1799.

Statement:

“All pure samples of a particular chemical compound contain similar elements combined in the same proportion by mass”.

Explanation:

The law is based on the fact that when elements combine to form a given compound, they do so in fixed proportions by mass so that all pure samples of that compound are identical in composition by mass.

Example: The law of definite proportion is practically proved by the example of Berzelius experiment.

BERZELIUS EXPERIMENT

Introduction:

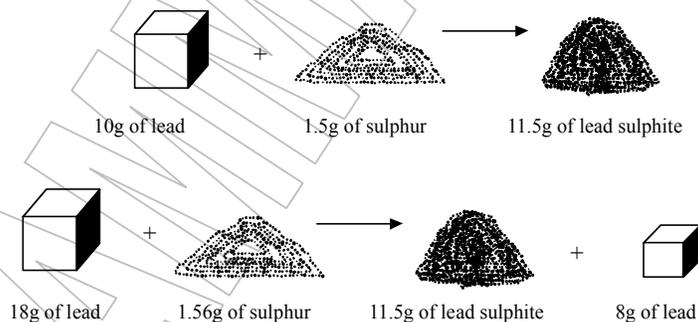
Swedish chemist J.J Berzelius was demonstrates the validity of the statement of this law.

Detail of experiment:

In this experiment, Berzelius heated 10 gm of lead (Pb) with various amount of sulphur (S). He got exactly 11.56 gm of lead sulphide (PbS). And excess of sulphur was left over, when he used 18 gm of lead (Pb) with 1.56 gm of sulphur, he got exactly 11.56 gm of PbS and 8 gm of lead remained unused.



These reactions are shown diagrammatically in figure.



Berzelius's experiment illustrates the law of definite proportions.

(iii) THE LAW OF MULTIPLE PROPORTIONS:

Introduction:

The fact that the same element, can combine in more than one ratio to form different compounds was published by, John Dalton in 1803 in the form of this law.

Statement:

“The masses of one element that combine with a fixed mass of the other element are in the ratio of small whole numbers or simple multiple ratio.”

Example: The excellent illustration of law of multiple proportion is furnished, when the element carbon (C) and oxygen (O) combine together to form their oxides namely carbon monoxide (CO) and carbon dioxide (CO₂).

COMPOUND	MASS OF CARBON	MASS OF OXYGEN	RATIO OF OXYGEN
Carbon monoxide (CO)	12	16	1
Carbon dioxide (CO ₂)	12	32	2

(iv) **THE LAW OF RECIPROCAL PROPORTION:**

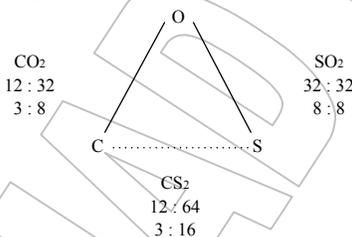
Introduction:

This is the fourth law of chemical combination and it put forward by a well-known chemist Richter in 1792-94.

Statement:

“When two different elements separately combine with the fixed mass of the third element, the proportion in which they combine with one another shall be either in the same ratio or same simple multiple of it”.

Example: The illustration of law of reciprocal proportion is provided when, 12 gm of carbon (C) combine with 32 gm of oxygen (O) to form carbon dioxide (CO₂) and 32 gm of sulphur (S) combine with same (fixed) mass of oxygen (O) i.e. 32 gm to form sulphur dioxide (SO₂).



The above example shows that the mass of C and S that combine with the same mass of O are in the proportion of 12:64 i.e. 3:16.

ATOM:

“The smallest particle of matter is called an atom.”

Example:

- Atom of Sodium (Na)
- Atom of Sulphur (s)

MOLECULAR:

“The smallest particle of matter formed by the combination of two or more like or unlike atoms is called a molecule.”

Example:

- Molecule of Nitrogen (N₂)
- Molecule of Water (H₂O)

FORMULA:

“The representation of molecule of a substance in terms of symbols is called formula.”

Example:

- Formula of Ammonia gas = NH₃
- Formula of Methane gas = CH₄
- Formula of Nitric acid = HNO₃

Types of Formula:

There are two types of formulae.

- (1) Empirical Formula
- (2) Molecular Formula

EMPIRICAL FORMULA:

“That formula which gives simple ratio between the atoms of different elements is called empirical formula.”

Example:

- Empirical formula of glucose = CH_2O
- Empirical formula of benzene = CH

MOLECULAR FORMULA:

“That formula which gives actual ratio between the atoms of different elements is called molecular formula.”

Example:

- Molecular formula of glucose = $\text{C}_6\text{H}_{12}\text{O}_6$
- Molecular formula of benzene = C_6H_6

Mathematically:

$$\eta = \frac{\text{M.F. weight}}{\text{E.F. weight}}$$

Where, η = The whole numbers ($n = 1, 2, 3, \dots$)

CHEMICAL REACTION:

“A process in which chemical change in the nature and composition of some substance (s) occurs is called a chemical reaction.”

Example:

- $\text{Fe} + \text{S} \rightarrow \text{FeS}$
- $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$

Types of Chemical Reactions:

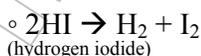
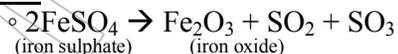
The some important and common types of chemical reaction are given below:

- (i) Decomposition reactions
- (ii) Addition Reaction or Combination Reaction
- (iii) Single displacement Reaction
- (iv) Double displacement Reaction
- (v) Combustion Reaction

(i) DECOMPOSITION REACTION:

“The reaction in which a chemical compound splits up in to two or more simple substance, when heat energy is given is called a decomposition reaction.”

Example:



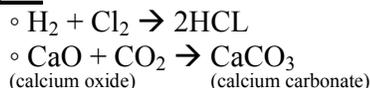
(ii) ADDITION REACTION:

“A reaction in which one single compound is formed by the combination of two or more substances is called addition reaction.”

Important Point or Remember:

- (a) These reactions are reverse of decomposition reactions.
- (b) These reactions are also called “Synthesis Reaction”.

Example:

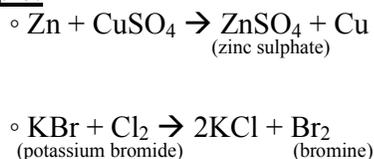


(iii) DISPLACEMENT REACTION:

“A reaction in which one atom or group of atoms of a compound is replaced by another action or group of atoms is called displacement reaction.”

These reactions are also called Single replacement reaction.

Example:

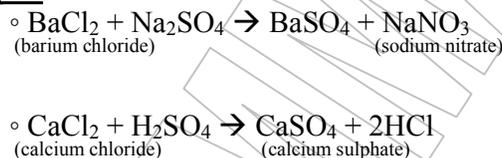


(iv) DOUBLE DISPLACEMENT REACTION:

“A type of reaction in which the two reactants are both decomposed to form new substances by excellent-changing their radicals.”

i.e. $AB + CD \rightarrow AD + CB$

Example:



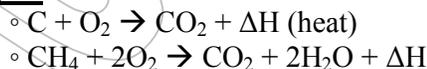
Another classes of Double Displacement Reaction are

- (a) Neutralization Reaction
- (b) Hydrolysis Reaction

(v) COMBUSTION REACTION:

“A reaction in which substances react with either free oxygen or oxygen of the air, with the rapid release of heat and flame, is called combustion reaction.”

Example:



CHEMICAL EQUATION:

“A chemical equation is a symbolic representation of chemical change in terms of symbols and formulae.”

Parts of Chemical Equation:

There are two parts of a chemical equation.

- (i) Reactants (ii) Products

REACTANTS:

“The substances which react together in a chemical reaction are called reactants.”

PRODUCTS:

“The substances which are form in a chemical reaction are called products.”

ATOMIC MASS:

“The average relative mass of the isotope of that element referred to the atomic mass of carbon which is taken as 12.0 a.m.u is called atomic mass.”

Example:

- Atomic mass of Nitrogen (N) = 7 a.m.u.
- Atomic mass of Hydrogen (H) = 1 a.m.u.

MOLECULAR FORMULA MASS:

“The sum of the atomic masses of all atoms present in the molecular formula of a substance or molecule is called molecular mass.”

Example:

- Molecular mass of Ammonia gas (NH₃) = 17 a.m.u.
- Molecular mass of Water (H₂O) = 18 a.m.u.

FORMULA MASS:

“The sum of the atomic masses of all atoms in a formula unit of the substance is called formula mass.”

Example:

- Formula mass of sodium chloride (NaCl)
Na = 23 x 1 = 23 a.m.u.
Cl = 35.5 x 1 = 35.5 a.m.u.
Formula mass of NaCl = 58.5 a.m.u.

MOLAR MASS:

“The relative molecular mass expressed in grams, is called molar mass of a substance.”

Example:

Molar mass of ammonia (NH₃)
N = 1 x 14 gm = 14 gm
H = 3 x 1 gm = 3 gm
Molar mass of NH₃ = 17 gm

MOLE:

“The atomic mass, molecular mass or formula mass of a substance expressed in grams is equal to 1 mole of that substance.”

Example:

- 1 mole of Nitrogen molecule (N₂) = 14 gm
- 1 mole of Iron oxide (Fe₂O₃) = 160 gm
- 1 mole of carbon atom (C) = 12 gm

Formula:

Mole is a quantity as well as a number.

$$\circ \text{ Number of moles} = \frac{\text{Given mass of substance in gm}}{\text{Atomic / Molecular / Formula mass}}$$

$$n = \frac{m}{M}$$

Unit of Mole:

In SI system, the unit of mole is "mol".

AVOGADRO'S NUMBER (N):

"The number of particles (atoms or molecules) in one mole of a substance is 6.023×10^{23} . This is known as Avogadro's Number."

Example:

- 1 mole of C = 12 gm = 6.023×10^{23} atoms of carbon
- 1 mole of H₂O = 18 gm = 6.023×10^{23} molecules of water
- 1 mole of Na = 23 gm = 6.023×10^{23} ions of Na⁺

Uses of Avogadro's Number:

- (a) To calculate the mass of single atom or molecule of any substance.
- (b) To calculate the number of atoms or molecules in a given sample of substance.

Formula:**(i) In Terms of Mass:**

$$\circ \text{ Number of atoms / molecules} = \frac{\text{Mass of substance}}{\text{Atomic / Molecular mass}} \times \text{Avogadro's number}$$

$$\text{no. of atoms / molecules} = \frac{m}{M} \times N_A$$

(ii) In Terms of Mole:

$$\circ \text{ Number of atoms / molecules} = \text{Mole number} \times \text{Avogadro's number}$$

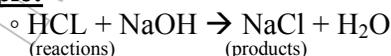
$$\text{no. of atoms / molecules} = n \times N_A$$

(iii) For Mass of One Atom:

$$\circ \text{ Mass of one atom} = \frac{\text{Atomic mass in gram}}{\text{Avogadro's number}}$$

(iv) For Mass of One Molecule:

$$\circ \text{ Mass of one molecule} = \frac{\text{Molecular mass in gram}}{\text{Avogadro's number}}$$

Example:

Important Point or Remember:

- (a) The reactants and products are separated from one another by using the sign arrow (\rightarrow)
- (b) The single arrow (\rightarrow) represents one way reactions.
- (c) The double arrow (\rightleftharpoons) represents two way reactions.

DIFFERENCE BETWEEN:

EMPIRICAL FORMULA	MOLECULAR FORMULA
(a) A formula that gives simple ratio is called empirical formula.	A formula that gives actual ratio is called molecular formula.
(b) It gives only the relative number of each type of atoms present in a molecule.	It gives only the actual number and type of atoms in a molecule.

Problems

- (a) Calculate the molecular mass of the following:
(i) $C_2H_5 - OH$ (ii) NH_3 (iii) CH_4 (iv) NO_2
- (b) Calculate the formula mass of the following:
(i) H_2SO_4 (ii) $H_2S_2O_7$ (iii) Fe_2O_3 (iv) $BaCl_2$
- (c) Calculate number of moles in:
(i) 32 gm of HNO_3
(ii) 117 gm of $CH_3COOC_2H_5$
(iii) 79 gm of $NaOH$
- (d) Calculate:
(i) 72 gm of C_6H_6
(ii) 206 gm of $C_{12}H_{22}O_{11}$
(iii) 4 gm of S
(iv) 1.5 gm of O
- (e) Calculate number of molecules in:
(i) 8 gm of SO_3
(ii) 8.8 gm of CO_3
(iii) 31 gm of H_2O
(iv) 23.7 gm of CH_4
- (f) Calculate the mass of one molecule of:
(a) $C_3H_7 - OH$ (ii) HNO_3 (iii) Cl_2 (iv) H_2O_2
- (g) Potassium chloride ($KClO_3$) is decomposed on heating gives Potassium chloride (KCl) and oxygen (O_2)
◦ $2KClO_3 \rightarrow 2KCl + 3O_2$
Calculate the amount of KCl by heating 7.2 moles of $KClO_3$.
- (h) At high temperature sulphur combines with iron (Fe) to form brown-black iron sulphide (FeS).
◦ $Fe + S \rightarrow FeS$
- (i) Sulphur dioxide (SO_2) oxidizes in a chamber in to sulphur trioxide (SO_3) at high temperature.
◦ $2SO_2 + 3O_2 \rightarrow 2SO_3$
Calculate the amount of SO_3 formed by 211 moles of Oxygen (O).
- (j) Ammonia gas is formed (NH_3) by the combination of nitrogen (N_2) with hydrogen (H_2).
◦ $N_2 + 3H_2 \rightarrow 2NH_3$
Calculate the amount of NH_3 gas formed by 13 moles N_2 is used.

Balance the following Equations

1. $C_2H_4 + O_2 \rightarrow CO_2 + H_2O$
2. $2CaCO_3 \rightarrow CaO + CO_2$
3. $Fe + O_2 \rightarrow Fe_2O_3$
4. $KClO_3 \rightarrow KCl + O_2$
5. $FeSO_4 \rightarrow Fe_2O_3 + SO_2 + SO_3$
6. $KClO_3 \rightarrow KCl + O_2$
7. $Na + Cl_2 \rightarrow NaCl$
8. $CaCO_3 + HCl \rightarrow CaCl_2 + H_2O + CO_2$
9. $MnO_2 + HCl \rightarrow MnCl_2 + H_2O + CO_2$
10. $CH_4 + O_2 \rightarrow CO_2 + H_2O + \text{heat}$
11. $ZnCl_2 + AgNO_3 \rightarrow Zn(NO_3)_2 + AgCl$
12. $Cu + H_2SO_4 (\text{conc.}) \rightarrow CuSO_4 + H_2O + SO_2$
13. $NH_4Cl + Ca(OH)_2 \rightarrow CaCl_2 + H_2O + NH_3$
14. $Al + Cl_2 \rightarrow AlCl_3$
15. $NaBr + Cl_2 \rightarrow NaCl + Br_2$
16. $H_2O_2 \rightarrow H_2O + O_2$
17. $FeSO_4 \rightarrow Fe_2O_3 + SO_2 + SO_3$
18. $H_2S + O_2 \rightarrow SO_2 + H_2O$
19. $Zn + HNO_3 \rightarrow Zn(NO_3)_2 + H_2O + NH_4NO_3$
20. $PCl_3 + H_2O \rightarrow H_3PO_3 + HCl$
21. $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
22. $FeCl_3 + H_2O \rightarrow Fe_2O_3 + HCl$
23. $Na + H_2O \rightarrow NaOH + H_2$
24. $Pb(NO_3)_2 + HCl \rightarrow PbCl_2 + NH_3$
25. $Mg + N_2 \rightarrow Mg_3N_2$
26. $Fe + HCl (\text{dil}) \rightarrow FeCl_2 + H_2$
27. $Pb(NO_3)_2 \rightarrow PbO + NO_2 + O_2$
28. $Na_2CO_3 + HClO_3 \rightarrow NaClO_3 + CO_2 + H_2O$
29. $NH_4OH + H_2SO_4 \rightarrow (NH_4)_2SO_4 + H_2O$
30. $Al + O_2 \rightarrow Al_2O_3$
31. $ZnO + NaOH \rightarrow Na_2ZnO_2 + H_2O$
32. $K_2O + H_2O \rightarrow KOH$

MUHAMMAD HASSAM

CHAPTER # 03

ATOMIC STRUCTURE

ATOM

Democritus Theory About Atom:

According to Greek Philosopher Democritus,
“All matter is composed of tiny, indivisible particles called atom.”

Modern Concept About Atom:

Modern concept about atom is that,
“The smallest particle of an element, which is found in nature is called atom.”

FUNDAMENTAL PARTICLES OF ATOM

According to modern concept, an atom is made up of numerous sub-atomic particles. Out of these, three sub-atomic particles have fundamental importance:

1. Electron
2. Proton
3. Neutron

1. CHARACTERISTICS OF ELECTRON:

- (a) It is negatively charged particle.
- (b) Its mass is equal to 9.11×10^{-31} kg or 5.48×10^{-4} a.m.u.
- (c) The charge on electron is equal to 1.6×10^{-19} Coulomb.
- (d) Electrons are present around the nucleus of an atom.

2. CHARACTERISTICS OF PROTON:

- (a) It is positively charged particle.
- (b) Its mass is equal to 1.67×10^{-27} kg or 1.007 a.m.u.
- (c) It is 1837 times heavier than an electron.
- (d) The charge on proton is equal to 1.602×10^{-19} Coulomb.
- (e) Protons are present in the nucleus of an atom.

3. CHARACTERISTICS OF NEUTRON:

- (a) Neutron has no charge.
- (b) Its mass is equal to 1.67×10^{-27} kg or 1.008 a.m.u.
- (c) It is 1842 times heavier than an electron.
- (d) Neutrons are also present in the nucleus of an atom.

DALTON'S ATOMIC THEORY

Introduction:

In 1808, an English school teacher John Dalton presented a scientific theory on atom.

The Fundamental Postulates of Dalton's Theory:

The main postulates of Dalton's atomic theory are as follows:

- (a) Matter is composed of tiny indivisible particles called atoms.
- (b) Atoms can neither be created nor be destroyed.
- (c) Atoms of the same element are similar but different from the atoms of other elements.
- (d) When atoms combine with other atoms they do so in simple ratio.
- (e) Chemical combination takes place between small whole number of atoms and all chemical changes produced are result of it.

Drawbacks in Dalton's Atomic Theory:

According to the latest research on the structure of atom and matter, the following defects are observed in Dalton's atomic theory.

- (a) Atom is not indivisible. It consists of charged particles called electron and proton and neutron particles called neutron.
- (b) The atoms of the same element may differ in their atomic masses and such atoms of an element are called isotopes.
- (c) The ratio between atoms of various elements in many molecules of the organic compounds is not constant.
- (d) Matter can be converted into energy by Einstein's equation. $E = mc^2$

KEY POINT:

Isotopes:

Elements that have same atomic number but different atomic mass.

◦ C = Speed of light.

i.e. 3.0×10^8 m/s.

DISCOVERY OF ELECTRON:

Electron was discovered by an experiment, called "Discharge Tube Experiment".

This experiment was performed by scientists named, "Crooks", "Faraday" and "J.J. Thomson".

DISCHARGE TUBE EXPERIMENT:

In this experiment, a simple glass discharge tube is constructed by fusing two metallic plates, (Electrode) through a glass tube at both ends. A very high voltage current was passed through gases at very low pressure (1 mm of Hg), the rays begin to originate from Cathode, which are called "Cathode rays". Cathode rays are negatively charged particles called "Electrons".

PROPERTIES OF CATHODE RAYS:

On the basis of experimental observations, the Cathode rays possess the following properties.

- (a) Cathode rays travel in straight lines away from the cathode.
- (b) Cathode rays carry negative charge.
- (c) The rays emerging normally from the cathode.
- (d) The rays, upon striking glass, cause this material to glow.
- (e) These rays consist of particles called Electron carrying a fixed unit of charge and a fixed mass.
- (f) The cathode rays are easily deflected by a magnetic field.

DISCOVERY OF PROTON:

A well known scientist, Goldstain discovered positive rays other than cathode rays in a gas discharge tube experiments.

In this experiment, Goldstain observed that these positive rays were found traveling in the opposite direction to that of the cathode rays. He named these positive rays, as Proton.

DISCOVERY OF NEUTRON:

In 1932, a well known scientist James Chadwick discovered another fundamental particle of atom by an experiment. He named it "Neutron".

ATOMIC NUMBER:

"The number of proton present in the nucleus of an atom, represent the atomic number of an element."

Symbolic Representation:

The atomic number is represented by "Z".

KEY POINTS:

Atom is electricity neutral because the number of protons and electrons are equal in an atom.

ATOMIC MASS:

“The sum of the protons and neutrons in the nucleus of an atom is called atomic mass.”

It may also be called “Mass number”.

Symbolic Representation:

The mass number is represented by “A”.

KEY POINTS:

Anion:

If an atom gain an extra electron & get –ve charge & because a –ve ion called Anion.

Cation:

If an atom lose an electron & get +ve charge & because a +ve ion called cation.

RUTHERFORD’S ATOMIC MODEL

Introduction:

In 1911, a well known scientist Rutherford carried out an experiment for the determination of the structure of atom.

Rutherford’s Experiment:

In this experiment, Rutherford took gold foil and a photographic plate, which was placed behind the foil. He bombarded “alpha-(α) particles on gold foil. From photographic plate he observed the extent of scattering of alpha (α) particles. From the flashes on the photographic plate, it was found that nearly all the particles passed through the foil with little or no deflection. However a few α – particles were deflected through large angles during bombardment and few bounced back from the gold foil.

Conclusions Drawn by Rutherford:

From this experiment Rutherford concluded that,

- (a) The major portion of the atom is empty.
- (b) The total mass of the atom is present in the nucleus.
- (c) The electrons revolve around the nucleus.
- (d) The positive charge is concentrated in the nucleus.
- (e) The size of the nucleus is very small as compared with the size of the atom.

KEY POINTS:

Alpha Particles:

A-particles are fast moving Helium nuclei.

Nucleus:

Nucleus is a central part of an atom which contains protons and neutrons.

The whole mass of an atom present in the nucleus.

Drawbacks in Rutherford’s Atomic Model:

Following drawbacks are found in Rutherford’s model,

- (a) According to classical electromagnetic theory, the revolving electron will lose energy continuously and the radius of the orbit would become smaller until the electrons fall into the nucleus. Hence atom should be unstable.
- (b) If the electrons emit energy continuously, they should form a continuous spectrum but actually line spectrum is obtained.

Thus according to classical principles, the Rutherford model of the atom could not exist.

BOHR'S ATOMIC MODEL

Introduction:

In 1913, a Danish physicist Niels Bohr proposed a new model for the structure of the atom.

Salient Features of Bohr's Atomic Model:

- Electron in an atom revolves around the nucleus in fixed circular path which he called orbits, or energy levels.
- When electron revolving in a particular orbit, it neither absorbs nor emits energy.
- If the electron absorbs energy equals to the energy difference between the two orbits, it jumps to higher energy level, away from nucleus.
- When electron loses energy, it returns to lower energy level, closer to the nucleus.
- If this energy is absorbed or emitted as light, a single quantum (photon) of absorbed or emitted light must account for the required energy difference, so that

$$E_2 - E_1 = \Delta E$$

$$E_2 - E_1 = h\nu$$

$$\Delta E = h\nu$$

Where,

E_1 = Energy of the electron in 1st orbit

E_2 = Energy of the electron in 2nd orbit

ΔE = Energy difference between two levels

ν = Frequency of the emitted light

h = Planck's Constant (its value = 6.625×10^{-34} J.s.)

- The angular momentum (mvr) of an electron in any orbit is integral multiple of $h/2\pi$.

$$mvr = \frac{nh}{2\pi}$$

Where,

m = mass of electron ($m_e = 9.11 \times 10^{-31}$ kg)

v = velocity of electron

r = radius of orbit

n = quantum numbers ($n = 1, 2, 3, \dots$)

π = Constant value = 3.142

h = Planck's Constant ($h = 6.625 \times 10^{-34}$ J.s.)

Conclusions Taken From Bohr's Theory

OR

Applications of Bohr's Theory

The conclusions taken from Bohr's atomic theory are given below:

- This theory gives clear explanation about the line spectrum of the elements.
- Bohr's theory gives idea about the stability of the atom.
- According to Bohr's theory, the number of electrons that can revolve in each orbit is limited.
Maximum numbers of electrons in particular orbit can be calculated by the formula

$$\text{Number of electron} = 2n^2$$

Where,

n = number of energy level or orbit.

The maximum no. of electron for the first four orbits

ENERGY LEVEL OR ORBIT	MAXIMUM NO. OF ELECTRONS
1 st = K shell	2
2 nd = L shell	8
3 rd = M shell	18
4 th = N shell	32

- (d) It also helps to explain the periodicity of the chemical properties of the elements.
(e) Ionization energy of the atom can be explained successfully with help of Bohr's theory.

ELECTRONIC CONFIGURATION:

“The arrangement of the electron around the nucleus of an atom in different shells (energy levels) is called electronic configuration.”

The distribution of electrons in various orbits, according to Bohr's theory, is shown in Table for the first 20 elements.

The Electronic Configuration of the first 20 elements

Period	Element	Symbol	At. No. (Z)	Electronic arrangement in various energy levels (orbits)				
				1 st	2 nd	3 rd	4 th	
1 st	Hydrogen	H	1	1				
	Helium	He	2	2				
2 nd	Lithium	Li	3	2	1			
	Beryllium	Be	4	2	2			
	Boron	B	5	2	3			
	Carbon	C	6	2	4			
	Nitrogen	N	7	2	5			
	Oxygen	O	8	2	6			
	Fluorine	F	9	2	7			
	Neon	Ne	10	2	8			
	3 rd	Sodium	Na	11	2	8	1	
		Magnesium	Mg	12	2	8	2	
Aluminium		Al	13	2	8	3		
Silicon		Si	14	2	8	4		
Phosphorous		P	15	2	8	5		
Sulphur		S	16	2	8	6		
Chlorine		Cl	17	2	8	7		
Argon		Ar	18	2	9	8		
4 th		Potassium	K	19	2	8	8	1
		Calcium	Ca	20	2	8	8	2

Characteristics of an Atom:

Following are some important characteristics of an atom across the “Periodic Table”.

ATOMIC RADIUS:

“Half of the distance between nuclei of two similar adjacent atoms is called atomic radius.”

Unit: Atomic radius generally expressed in Angstrom units, (A°), ($1 A^\circ = 10^{-10} m$).

Trend of Atomic Radius in Periodic Table:

- As in period the atomic number increases from left to right, the atomic radius decreases. Due to increase in nuclear charge on each atom.

The Atomic Radii of the Elements of 2nd Period

ELEMENTS	Na	Al	P	Cl	Ar
ATOMIC RADIUS (A ^o)	1.57	1.25	1.10	0.99	0.91

◦ When we move vertically down in a group. the atomic radius increases because of the addition of an extra orbit of electron in each atom.

The Atomic Radii of the Elements of Group I-A

ELEMENTS	Li	Na	K	Rb	Cs
ATOMIC RADIUS (A ^o)	1.23	1.57	2.03	2.16	2.35

IONIZATION ENERGY:

“The minimum amount of energy required to remove electrons from an isolated gaseous atom or ion in its ground state is called ionization energy.” (I.E.)

Unit: Ionization energy generally measured in electron volts. (eV). (1 eV = 1.6 x 10⁻¹⁹ J)

Trend of Ionization Energy in Periodic Table:

◦ As in period, the ionization energy of atoms increases from left to right. This is because, atoms have higher nuclear charge and a smaller radius.

Ionization Energies of 3rd Period Elements

ELEMENTS	Na	Mg	Si	P	S
IONIZATION ENERGY (eV)	5.14	7.64	8.15	10.49	12.97

◦ Ionization energies of atoms decreases down the group due to increase in size of atom and decrease in nuclear attraction.

Ionization Energies of I-A Group Elements

ELEMENTS	Li	Na	K	Rb	Cs
IONIZATION ENERGY (eV)	5.39	5.14	4.34	4.18	3.89

KEY POINTS:

Ionization:

When a neutral atom loses an electron it becomes a +velocity ion and the process is called ionization.

As ionization energy increases reactivity atoms decreases.

I.E depends upon the nuclear charge and the size of the atom.

ELECTRONEGATIVITY:

“The electronegativity of an atom is a relative measure of its tendency to attract a shared pair of electrons towards it self.”

Trend of Ionization Energy in Periodic Table:

◦ Electronegativity increases as we move from left to right in a period because of increasing nuclear charge and decreasing atomic radius.

◦ Halogens are most electronegative element in the periodic table Flourine (F) has the highest value = 4

Electronegativity value of 2nd Period Elements

ELEMENTS	B	C	N	O	F
IONIZATION ENERGY (eV)	2.0	2.5	3.0	3.5	4.0

◦ As we move downward in a group, the value of electronegativities of the elements decreases due to increase in atomic size.

Electronegativity value of the Halogens (VIII-A)

ELEMENTS	F	Cl	Br	I	At
IONIZATION ENERGY (eV)	4.0	3.0	2.8	2.5	4.0

Group VIII-A (Noble Gases) have zero value of Electronegativity because their orbits are complete.

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CHAPTER # 04

THE PERIODIC TABLE

PERIODIC TABLE:

“A chart or table in which element are arranged in order of increasing atomic number and also distributed into groups and periodic is known as Periodic table.”

CLASSIFICATION BY SCIENTISTS

AL-RAZI'S CLASSIFICATION:

In 910, Al-Razi was the first muslim scientist who classify the elements on that time on the basis of difference of physical and chemical properties. He also distinguished between metals and non-metals.

DOBEREINER'S TRIADS:

In 1829, a German chemist, J.W Dobereiner discovered that,
“With in a group of elements closely related in chemical properties, the atomic mass of the middle element is approximately the arithmetic mean of other two elements.”
These are called Dobereiner's triads.

TRIADS

	Elements	Atomic Mass	Arithmetic Mean
Triads	{ Lithium Sodium Potassium	7 23 39	$\frac{7 + 39}{2} = 39$
Triads	{ Chlorine Bromine Iodine	35.5 80 126.5	$\frac{35.5 + 126.5}{2} = 81$
Triads	{ Calcium Strontium Barium	40 87 137	$\frac{137 + 40}{2} = 88$

NEWLAND'S OCTAVES:

In 1866, an English chemist, Newland reported his law of octaves.
“The properties of every eight elements were similar to those of the first. This regularity of elements called as the Law of Octaves.”

Newland's Arrangement of Elements in Octaves

	No.												
H	1	Li	2	Be	3	B	4	C	5	N	6	O	7
F	8	Na	9	Mg	10	Al	11	Si	12	P	13	S	14
Cl	15	K	16	Ca	17	Cr	18	Ti	19	Mn	20	Fe	21

MENDELEEV'S PERIODIC TABLE:

In 1869, a Russian chemist Mendeleev designed his periodic table of elements. On the basis of physical and chemical properties he discovered a law known as Periodic law.
“The properties of element are the periodic functions of their atomic masses.”

Advantages of Mendeleev's Table:

Mendeleev's table offered the following advantages in understanding the chemistry of the elements.

1. In this table, elements arranged in horizontal rows or period and vertical columns or group. These groups further divided into sub groups.
2. There was a regular gradation in the physical and chemical properties in a sub group.
3. Mendeleev left many vacant spaces in this periodic table for unknown elements and also predict their properties.

Mendeleev's Periodic Table

H = 1			Ti = 50	Zr = 90	? = 180
			V = 51	Nb = 94	Ta = 182
			Cr = 52	Mo = 96	W = 186
			Mn = 55	Rh = 104.4	Pt = 197.4
			Fe = 56	Ru = 104.4	Ir = 798
			Ni=Co = 59	Pd = 106.6	Os = 199
			Cu = 63.4	Ag = 108	Hg = 200
		Be = 9.4	Mg = 24	Zn = 65.2	Cd = 112
		B = 11	Al = 27.4	? = 68	Ur = 116
		C = 12	Si = 28	? 70	Sn = 118
	N = 14	P = 31	As = 75	Sb = 122	
	O = 16	S = 32	Sc = 79.4	Te = 128.7	
	F = 19	Cl = 35.5	Br = 80	I = 127	
	Na = 23	K = 39	Rb = 85.4	Cs = 133	
Li = 7		Ca = 40	Sr = 87.6	Ba = 137	
		? = 45	Ce = 92		
		?Er = 56	La = 94		
		?Yt = 60	Di = 95		
		?In = 75.6	Th = 1187		
				Ti = 204	
				Pb = 207	

- The group number of an element indicates the highest valance state it can attain.
- Mendeleev's arrangement of elements helped to correct the atomic masses of elements.
- The first elements of groups I, II and III showed a similarity to the second elements of the next groups, i.e. groups II, III and IV respectively.

Drawbacks in Mendeleev's Table:

Inspite of many advantages, Mendeleev's periodic table of elements had certain drawbacks.

- Mendeleev's table does not give an idea about the structure of atoms.
- Mendeleev's table does not give any indication about the position of isotopes.
- The change in the atomic masses of two successive elements is not constant. Hence, it is impossible to predict the mass of missing element by knowing the atomic mass of two known elements.
- In certain place, elements which differ widely in properties are placed into the mass group. e.g. Alkali metals (Na, K etc) & Coinage metals (Cu, Ag etc).
- For placing elements in the proper groups, the order of the elements according to atomic mass was reversed in certain cases. e.g. Potassium (39) placed after Argon (40).

Conclusion:

From the above mentioned drawbacks it was concluded that the atomic mass of the element is not a fundamental property for classification.

MODERN PERIODIC LAW:

In 1913, a British physicist Moseley formed his periodic law.

“The properties of the elements are the periodic functions of their atomic number and not the atomic masses.”

Importance of Modern Periodic Law:

The modern periodic law provided logical and scientific ground for the classification of elements. This law also provided justification for correcting the position of some elements wrongly placed in Mendeleev's periodic table.

The elements are arranged in ascending order of their atomic number in the modern periodic table recommended by the International Union of Pure and Applied Chemistry (IUPAC), in 1957.

PERIODICITY:

“The repetition of physical and chemical properties of the elements after a regular interval is called periodicity.”

MODERN PERIODIC TABLE:

The modern periodic table consists of seven horizontal rows (periods) and eight vertical columns (groups).

PERIODS:

“Horizontal rows in the periodic table are called periods.”

There are seven periods in the modern periodic table.

1st Period:

First period consists of only two elements. Hydrogen (H^1) and Helium (He^2). It is called the shortest period.

2nd Period:

Second period consists of eight elements. From Lithium (Li^3) to Neon (Ne^{10}). It is called first short period.

3rd Period:

Third period also consists of eight elements. From Sodium (Na^{11}) to Argon (Ar^{18}). It is called second short period.

4th Period:

This period consists of eighteen elements. From Potassium (K^{19}) to Krypton (Kr^{36}). It is called first long period.

5th Period:

This period also consists of eighteen elements. From Rubidium (Rb^{37}) to Xenon (Xe^{54}). It is called second long period.

6th Period:

Sixth period consists of thirty-two elements. From Cesium (Ce^{55}) to Radon (Rn^{86}). It is called longest period.

Lanthanide series is included in this period.

7th Period:

Seventh period of the modern periodic table is incomplete period. This period starts from Francium (Fr^{87}).

Actinide series is included in this group.

GROUPS:

“Elements in vertical columns in the modern periodic table are called groups.”

The modern periodic table is divided into eight groups. But they are further divided into two sub-groups.

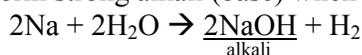
- (1) Sub group A
- (2) Sub group B

SUB GROUP A

Sub group A consists of eight groups. The normal elements of the earth crust are kept in sub group A.

Group I, The Alkali Metals:

1. Group I, starts from Lithium (Li³) to Francium (Fr⁸⁷) downward the group.
2. These elements are strongly electropositive.
3. They form ionic compounds.
4. Their Valency is +1.
5. They form strong alkali (base) when it reacts with water.



6. The alkali metals are light metals which are very reactive chemically.
7. Down the group ionization potential decreases with increase in atomic size.
8. They contain only an electron in their outer most shell.

KEY POINTS:

Hydrogen:

Hydrogen (H¹) is not included in any group because of their properties.

Ionization Potential:

It is the minimum amount of energy required to remove valence electron from an elements.
Radium (Ra¹⁸) is a radioactive elements.

Metalloid:

These are the elements that have both of properties of metal and non-metals.

Group II, The Alkaline-Earth Metals:

1. This group consists of Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) & Radium (Ra).
2. They form ionic compounds.
3. They contain two electrons in their valence shell. Therefore their valency is +2.
4. It is a group of moderately reaction metals, harder and less volatile than the alkali metals.
5. They are called "Alkaline Earth Metals" because they occur in nature as silicate minerals.
6. Their oxides and hydroxides are strongly basic.

Group III, The Boron or Aluminium Family:

1. This group consists of a metalloid (Boron) and also metals, these are Aluminium (Al), Gallium (Ga), Indium (In) & Thallium (Tl).
2. They are strongly electropositive.
3. They are quite reactive chemically.
4. They have three electron in their valence shell.
5. Therefore their valency becomes +3.
6. Due to small size, the value of ionization potential increases down the group.
7. There is increase in metallic character down the group.

Group IV, The Carbon & Silicon Family:

1. This group consists of five elements namely Carbon (C), Silicon (Si), Germanium (Ge), Tin (Sn) & Lead (Pb).
2. They usually form covalent bond except Tin and Lead.
3. They have four electrons in their outer most shell.
4. Electronegativity decreases down the group.
5. Their valency is +4.
6. Tin and Lead are electropositive and typical metals and they form ionic bonds.
7. The elements of this group evidently exhibit the intermediate character.

Group V, The Nitrogen Family:

1. This group consists of five elements. Nitrogen (N) and Phosphorus (P) are non-metals, Arsenic (As) and Antimony (Sb) are metalloids and Bismuth is weakly metallic.
2. They have five electrons in their valence orbit.
3. They usually form covalent bonds except Bismuth.
4. The tendency of forming covalent bond decreases down the group.
5. Bismuth only forms ionic bond, through the formation of (Bi^{3+}) cation.
6. This group displays a remarkable number of allotropes.
7. Only nitrogen is found to form cation (nitride).
8. The value of ionization potential is high due to small atomic size.

KEY POINTS:

Allotropes:

Elements which exist in two or more physics or molecular forms.

Electronegativity:

It is tendency to attracts shared pair of electron.

Group VI, The Oxygen Family:

1. This group consists of five elements. Oxygen (O) and Sulphur (S) are non-metals, Where as Tellurium (Te) and Selenium (Se) are metalloids and Polonium (Po) is considered as metal because of their basic properties.
2. In this group, Oxygen is a gas and other elements are solid.
3. The metallic character, ionic and basic nature increase regularly down the group, from oxygen to polonium.
4. The value of electronegativity of these elements are very high and decreases down the group.
5. In this group, all elements have six electrons in their valence orbit.
6. All elements of the group show the property of allotropy.

Group VII, The Halogen Family:

1. This group consists of fluorine (F) and Chlorine (Cl) as gases, Bromine (Br) as liquid, Iodine (I) and Astatine (At) as solid.
2. Halogens are highly electronegative elements and their electronegativity decreases down the group.
3. The halogen are very active non-metals.
4. They form covalent bonds.
5. They have seven electrons in their valence orbit. As their valency is -1.
6. All the halogen exist as diatomic molecules e.g. Cl_2 , Br_2 .
7. As halogens are extremely active so they not found in free state.

Group VIII, Zero Group OR Noble Gases:

1. They are all monoatomic gases.
2. This group consists of Helium (He), Neon (Ne), Argon (Ar), Krypton (Kr), Xenon (Xe) and Radon (Rn).
3. Their valence orbit is complete.
4. All noble gases except radon are normally present in the atmosphere.
5. They have low boiling point. E.g. Ne = -246°C .
6. These elements are mostly chemically unreactive.
7. These gases are mostly obtained from air by liquification process.

SUB GROUP B

Sub group B consists of eight group. Heavy metals are kept in this group. These elements are also called “Transition Elements”.

Properties & Characteristics of Transition Elements:

1. The all elements included in sub group B called Transition elements.
2. They all are heavy metals.
3. They show variable valencies.
4. They have remarkable ability to form complex ion by coordination.
5. Their compounds are formed by coordinate covalent bond.
6. Their compounds are usually coloured.
7. These elements have incomplete inner electron shells.
8. The bond between the atoms of these elements are very strong.
9. They have high melting point. e.g. Fe = 1535 °C
10. They have two rare-earth series of elements, “The Lanthanide” and “The Actinide” series.
11. Many compounds of these elements are paramagnetic, i.e. they are attracted into a magnetic field.

GROUP TRENDS:

The general group trends are given below:

1. Ionization Potential (I.P).

- (i) Down the group value of ionization potential decreases from top to bottom.
- (ii) In a period, its value increases from left to right.

2. Electronegativity:

- (i) Generally in a group, the value of electronegativity decreases from top to bottom.
- (ii) Its value increases from left to right in a period.

3. Metallic Character:

- (i) In a group, metallic character increases from top to bottom.
- (ii) Metallic character decreases in a period from left to right.

4. Ionic Size:

- (i) Ionic size increases down the group.
- (ii) In a period, ionic size decreases from left to right.

5. Acidic Character of Oxides:

- (i) In a group, acidic character of the oxides decreases from top to bottom.
- (ii) Acidic character of oxides increase from left to right in a period.

4. Basic Character of Oxides:

- (i) In a group, basic character of oxides increases from top to bottom.
- (ii) In a period, basic character of oxides decreases from left to right.

5. Atomic Sizes:

- (i) Atomic size increases down the group by increasing atomic number.
- (ii) Atomic size decreases from left to right in a period by increasing charge of an atom.

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CHAPTER # 05

CHEMICAL BONDING

CHEMICAL BONDING:

“The attractive force which bonds atoms together is called a chemical bond.”

Generally the atoms combine with another in three ways.

- (a) By losing the electrons
- (b) By gaining the electrons
- (c) By sharing the electrons

OCTET RULE:

“The tendency of atoms of elements to acquire an outermost shell of eight electrons is called octet rule.”

DUPLET RULE:

“The tendency of atom to acquire an outermost shell of only two electrons is called duplet rule.”

TYPES OF CHEMICAL BOND

There are two main types of chemical bonds.

1. Ionic (or Electrovalent) Bond.
2. Covalent Bonds.

Covalent bond is further sub divided in to:

- (a) Ordinary Covalent bonds.
- (b) Co-ordinate (or Dative) Covalent bond.

1. IONIC BOND

“The bond which is formed by losing or gaining of electrons is called ionic bond.”

Ionic bond is also called “Electrovalent bond”.

Ionization Energies & Ionic Bonding:

Those elements which have lowest value of ionization energy most readily form ionic bond by losing electrons.

- (a) The elements of group I-A & II-A have lowest value of I.E. so they form ionic bonds.
- (b) Most of transition metals also form ionic bonds in most of their compounds.
- (c) Metals of group III-A (Al group) and the heaviest elements in group IV-A (Tin & lead) also formed ionic bond due to lowest value of I.E.

Electronegativity & Ionic Bonding:

These elements which have highest value of Electronegativity most readily form ionic bond by gaining electrons.

- (a) Halogens (Group VII-A) have greatest electronegativities, atom of these elements readily form ionic bonds by accepting electrons.
- (b) In group VI-A Oxygen and Sulphur and in group V-A Nitrogen (N), also have relatively high electronegativities and tend to form ionic bonds, but less readily than halogens.

KEY POINTS:

Ionization Energy:

The minimum amount of energy required to remove electron from an atom.

Electronegativity:

The tendency of an atom to attract shared pair of electrons.

IONIC COMPOUND:

“Compound which are formed by ionic bonding are called ionic compounds.”

Example:

- Sodium Chloride (NaCl)
- Magnesium Oxide (MgO)

Properties of Ionic Compounds:

Some of the important properties of ionic compounds are enlisted below:

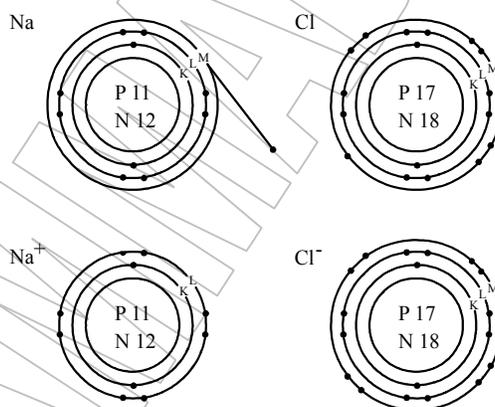
- Ionic compounds are generally hard solids, because of strong binding forces.
- They have high melting and boiling points.
e.g. M.P. of NaCl = 801 °C
- They are mostly inorganic compounds.
- They are usually insoluble in organic compounds and dissolve in water and other inorganic solvents.
- They dissolve in water to form electrolytes which readily conduct electricity.
- They do not conduct electricity in solid state.

EXAMPLES OF IONIC BOND

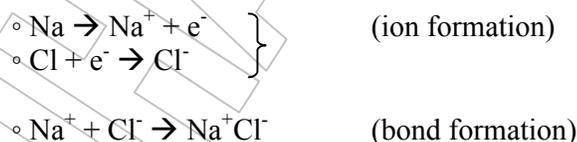
(a) Formation of Sodium Chloride (NaCl):

The reaction of sodium with chlorine to form sodium chloride is a good example of an ionic bond.

In this case sodium (Na) atom loses an electron which is gained by chloride (Cl) atom. As a result of which sodium cation (Na^+) and chloride anion (Cl^-) is formed and the outer most shell of both ions get eight electrons according to octet rule. These oppositely charged ions are held together by electrostatic force of attraction and ionic bond is formed.



Reactions:



(b) Formation of Magnesium Oxide (MgO):

The reaction of magnesium with oxygen to form magnesium oxide is another good example of formation of ionic bond.

In this case magnesium (Mg) atom loses two electrons which is gained by oxygen (O) atom. As a result of which magnesium cation (Mg^{2+}) and oxygen anion (O^{2-}) is formed and the outer most shell of both ions gets eight electrons according to octet rule. These oppositely charged ions are held together by electrostatic force of attraction and ionic bond is formed.

Reactions:**2. COVALENT BOND****Introduction:**

In 1916, the idea of the covalent bond was first introduced by a well known scientist named “G.N. Lewis”.

Statement:

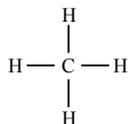
“A chemical bond which is formed by the mutual sharing of electrons is known as Covalent bond.”

Representation:

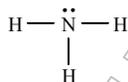
A covalent bond is generally represented by a short straight line between two bonded atoms.

Example:

(a) Methane (CH_4)



(b) Ammonia (NH_3)

**TYPES OF COVALENT BOND**

Generally there are three types of covalent bond, which is

- Single Covalent bond
- Double Covalent bond
- Triple Covalent bond

(a) Single Covalent Bond:

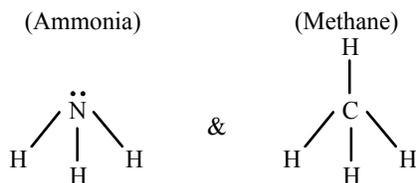
“A covalent bond formed by the mutual sharing of one electron pair is called a single covalent bond.”

Representation:

A single covalent bond is denoted by single short straight line (—).

Example:

- In $\text{H}-\text{Cl}$, $\text{H}-\text{H}$, $\text{Cl}-\text{Cl}$.
- In NH_3 three single covalent bond and in CH_4 four single covalent bonds are found



(b) Double Covalent Bond:

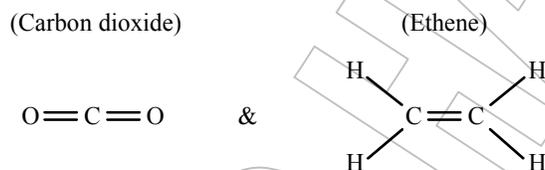
“A covalent bond formed by sharing two electron pairs, is called double covalent bond.”

Representation:

A double covalent bond is denoted by double short straight line (\equiv).

Example:

- In $\text{O}=\text{O}$.
- In Ethene (C_2H_4).
- In CO_2 two double covalent bonds are found.



(c) Triple Covalent Bond:

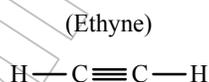
“A covalent bond is formed by sharing of three electron pairs is called triple covalent bond.”

Representation:

A triple covalent bond is denoted by triple short straight line (\equiv).

Example:

- In $\text{N}\equiv\text{N}$
- In ethyne (C_2H_2)



POLAR COVALENT BOND:

“A covalent bond formed between unlike atoms is known as polar covalent bond.”

OR

“A covalent bond in which the shared pair of electrons are attracted unequally by both atoms is called polar covalent bond.”

Example:

Bond between $\text{H}-\text{Cl}$, H_2O

NON-POLAR COVALENT BOND:

“A covalent bond is formed between like atoms is called non-polar covalent bond.”

Example:

Bond between CH_4 , O_2 , N_2 & H_2 etc.

Key Points:

- (a) A polar bond is formed between such two atoms which have electronegativity difference greater than 0.4.
e.g. H—Cl, H₂O
- (b) A non-polar bond is formed between such two atoms which have electronegativity difference less than 0.4
e.g. H—Cl, H₂O

POLAR COMPOUNDS:

“Such type of compounds which have polar covalent bond are called Polar compounds.”

Example:

Some example of polar compounds are,

- Hydrochloric acid (HCl)
- Water (H₂O)
- Ammonia (NH₃)

NON-POLAR COMPOUNDS:

“Such type of compounds which have non-polar covalent bond are called non-polar compounds.”

Example:

- Methane (CH₄)
 - Alcohol (C₂H₅—OH)
 - Petrol (C₆H₁₄)
- are good examples of non-polar compounds.

COVALENT COMPOUNDS:

“Compounds which are formed by covalent bonding (Polar or Non-Polar) are called covalent compounds.”

Example:

- Hydrochloric acid (HCl)
- Benzene (C₆H₆)
- Glucose (C₆H₁₂O₆)

Properties of Covalent Compounds:

Some of the important properties of covalent compounds are enlisted below:

- (a) Covalent compounds are mostly organic compounds.
- (b) Covalent compounds are found in all three state solid, liquid and gas at room temperature.
e.g. ◦ Wax is solid.
◦ Petrol is liquid.
◦ Methane & Carbon dioxide are gases.
- (c) They have low melting and boiling points.
e.g. M.P of Alcohol (C₂H₅—OH) = 78.5 °C
- (d) Covalent molecular compounds are generally soft because their molecules are held together by a weak force known as intermolecular force of attraction.
- (e) Covalent molecular compounds are generally soluble in organic solvents but are insoluble in water.
- (f) Non-polar covalent compounds do not conduct electricity in solid, molten or solution form.
- (g) Polar covalent compounds can conduct electricity in aqueous solution form.
- (h) Covalent compounds are mostly volatile.
e.g. ◦ Petrol (C₆H₁₄)
◦ Benzene (C₆H₆)

BOND POLARIZATION:

“The phenomenon of formation of such covalent bonds which carry partial positive and partial negative charges is called bond polarization. Such bond is called a “dipole”.

DIPOLE:

“The separation of charge in a bond or a molecule with a positively charged end and a negatively charged end, this is known as a dipole.”

DIPOLE-DIPOLE INTERACTION:

“The electrostatic attraction is created between the molecules, when partially negative end of one molecule is directed towards the partially positive end of the neighbouring molecules. This electrostatic attraction is known as dipole-dipole interaction.”

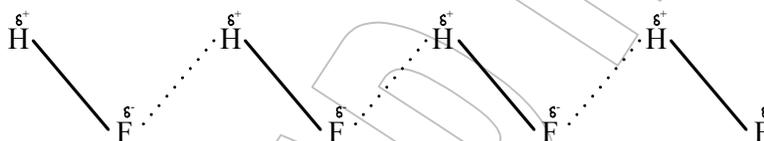
HYDROGEN BOND:

“An intermolecular force which arises when Hydrogen atom is bonded to a certain highly electronegative element (F, Cl, O) is called Hydrogen bonding.”

or it may also be define as,

“It is the name given to an intermolecular force which arises when Hydrogen atom bonded to electronegative atom (F, Cl, O) is weakly attracted by the electronegative atom of the neighbouring molecule.”

- Hydrogen bond is a very weak bond.

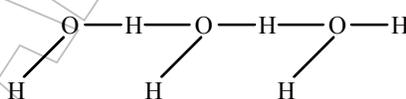


- “Hydrogen bonding in HF between H & F (Broken lines)”

Effect of Bond Polarization & Hydrogen Bonding on Physical Properties of Compounds

The physical properties of molecular substances are greatly affected by the force of attraction between the molecules. These forces depend on polarization of any bond in the molecule (ionic or covalent). The case of hydrogen bonds represents the strongest intermolecular attraction of all.

In fact, water (H₂O) is the most common example of a substance whose physical properties are strongly influenced by hydrogen bonding.



- Hydrogen bonding (Broken lines) in water”

The unusual properties of water which can be explained on the basis of hydrogen bonding are

(a) Melting & Boiling Points.

The melting and boiling point of water are considerably higher than the other hydrides of group VI-A such as H₂S. Because of hydrogen bonding in water, more energy is required to separate the molecules from each other.

Melting and Boiling Points of VI Group Hydrides

Substance	Molecular mass	Mp (°C)	Bp (°C)
H ₂ O	18.0	0 °C	100 °C
H ₂ S	34.1	83 °C	-16 °C
H ₂ Se	81	-66 °C	-41 °C
H ₂ Te	129.6	-49 °C	-2 °C

(b) Latent Heat of Fusion & Vapourization:

The values of latent heat of fusion & vapourization of water are relatively high as compared to other similar compounds due to hydrogen bonding.

- The value of latent heat of fusion of H₂O is 6 KJ/mole
- The value of latent heat of vapourization of H₂O is 41 KJ/mole, which are unusually a higher values.

(c) Specific Heat Capacity:

The value of specific heat capacity of H₂O is also very high as compared to that of other liquids due to presence of hydrogen bonding between the water molecules.

- Specific heat capacity value of H₂O is 4.2 J/g °C.

Specific Heat of Water Compared with that of some other Substances

Substance	Specific Hat (J/g °C)
Water	4.2
Ethanol	2.4
Iron	0.45

(d) Density:

The density of water in solid state is usually less than that liquid state. Because of hydrogen bonding when temperature goes down from 4 °C the water molecules show expansion. That's why its density becomes lower.

- The density of H₂O at 4 °C = 1 gm/cm³

(e) Solubility:

Water molecule has dipoles because of hydrogen bonding. These dipoles create polarization phenomenon. Because of that,

- All polar compounds dissolve (soluble) in water (ionic or covalent).
e.g. NaCl, HCl, NaOH, CuSO₄ etc.
- All non-polar compounds are not soluble in H₂O because they have no dipoles.
e.g. Benzene (C₆H₆), Petrol (C₆H₁₄)

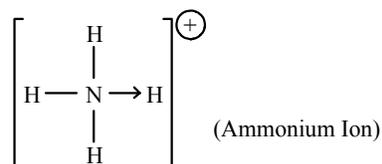
CO-ORDINATE COVALENT BOND:

“A bond in which the shared pair of electrons is denoted by one atom only, is called a Co-ordinate covalent bond.”

Representation:

A co-ordinate covalent bond is represented by an arrow (→). The head of the arrow directs towards the atom which doesn't contribute any electron.

i.e. From Donor to Acceptor



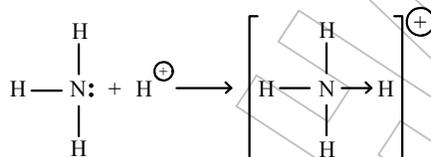
Examples:

Some common examples of co-ordinate covalent bond are given below:

(a) Formation of Ammonium Ion:

Ammonium ion is formed by Ammonia gas (NH_3). For this NH_3 acts as an electron pair donor or lewis base or ligands as the lone pair of electrons left on nitrogen can be donated to a lewis acid (electron pair acceptor) for a central metal atom.

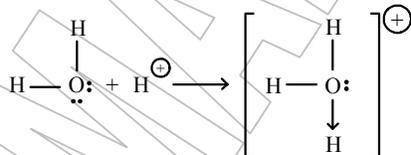
When NH_3 gas reacts with hydrogen ions (H^+) in aqueous solution of an acid an ammonium ion is formed.



(b) Formation of Hydroxonium Ion:

Water also acts as ligand.

“When water reacts with hydrogen ions (H^+) in aqueous solution of an acid a hydroxonium ion (H_3O^+) is formed. This ion may also be called hydronium ion.”



CO-ORDINATION COMPOUNDS:

“When a fixed number of molecules or ions are bonded to a metal atom through co-ordinate covalent bond, the compounds are formed called co-ordination compounds.”

It is also called “Complex Compounds”.

Example:

- Hydrated Copper sulphate ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$)
- Potassium Permanganate (KMnO_4)
- Potassium Dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$)
- Hemoglobin (compound of iron, the red constituent of blood)

LIGANDS:

“Group which surrounded by the central atom and can donate electron pair called ligands.”

Example:

- Water (H_2O)
- Ammonia (NH_3)
- Cyanide ion (CN^-)

CO-ORDINATION NUMBER:

“The number of bonds formed by ligands to the central metal atom or ion in a complex is called co-ordination number.”

Important Points to Remember:

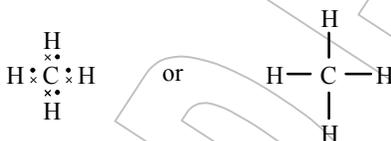
- Co-ordinate covalent bond is also called “Dative covalent bond”.
- All the transition metals of sub group-B in periodic table generally form co-ordination or complex compounds.
- The properties of co-ordination compounds are same as ordinary covalent compounds.
- Co-ordination compounds are usually coloured
e.g.
 - $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} \rightarrow$ blue in colour
 - $\text{KMnO}_4 \rightarrow$ violet (purple) in colour
 - $\text{K}_2\text{CrO}_4 \rightarrow$ yellow in colour

SHAPES OF COVALENT MOLECULES

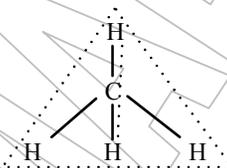
The shapes of some covalent molecules are given below.

(a) METHANE (CH_4):

As we know that the carbon atom has four electrons in its valence shell. In a molecule of methane (CH_4), it shares these with four hydrogen (H) atoms forming four covalent bonds.



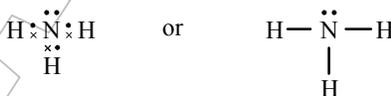
Electronic Structure:



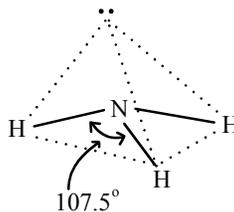
The molecule of methane has regular tetrahedral shape. The angle between any two C—H bonds is, about 109.5° .

(b) AMMONIA (NH_3):

As we know that the nitrogen atom has five electrons in its valence shell. In a molecule of ammonia (NH_3), it shares three electrons out of five, with three hydrogen atoms (H) forming three covalent bonds and leaving a lone pair of electrons.



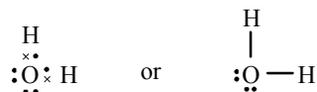
Electronic Structure:



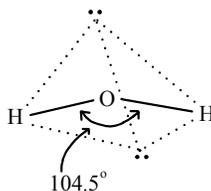
The molecule of ammonia (NH_3) has trigonal pyramidal shape. The angle between N—H is, about 107.5° .

(c) **WATER (H₂O):**

As we know that the oxygen atom has six electrons in its valence shell. In a molecule of water (H₂O), it shares only two electrons out of five, with two hydrogen atoms (H) forming two covalent bonds and leaving two lone pairs of electrons.



Electronic Structure:



The molecule of water (H₂O) has regular tetrahedron shape. The lone pairs (") have a remarkable tendency to repel shared pairs. Consequently the bond angle between H — O reduced from 109.5° to 104.5°.

DIFFERENCE BETWEEN

IONIC BOND	COVALENT BOND
(a) Ionic bond is formed by gaining or losing electrons between atoms.	Covalent bond is formed by sharing electrons between atoms.
(b) Ionic bond is formed when the electronegativity difference of two atoms is more than 1.7.	Covalent bond is formed if the electronegativity difference of two atoms is less than 1.7.
(c) It is a strong bond.	It is a weaker bond than ionic bond.

Key Points:

Volatile substances are those which can evaporate at room temperature. e.g. spirit, petrol etc.

IONIC COMPOUNDS	COVALENT COMPOUNDS
(a) Compound which are formed by ionic bond are called ionic compounds.	Compounds which are formed by covalent bonding are called covalent compounds.
(b) They are generally solid at room temperature.	They exist in all three states. i.e. solid, liquid & gas.
(c) They are good conductor of electricity in fused state.	They are bad conductors of electricity in purest form.
(d) They are non-volatile.	They are mostly volatile.
(e) They have high melting and boiling points. e.g. M.P. of NaCl = 801°C.	They have relatively low melting and boiling points e.g. B.P. of Benzene = 80°C.
(f) They are soluble in polar solvents.	They are soluble in non-polar solvents.
(g) They are usually inorganic compounds.	They are usually organic compounds.
(h) They are hard and brittle.	They are usually soft.

CHAPTER # 06

STATES OF MATTER

MATTER:

“Any thing that has mass and occupies space is called matter.”

STATES OF MATTER

There are three common states of matter. i.e.

- (a) Solid
- (b) Liquid
- (c) Gas

(a) SOLID STATE:

- (i) A solid possesses both definite shape and definite volume.
- (ii) In solid state molecules are tightly packed.
- (iii) It is not compressed on applying high pressure because their molecules are very close to each other.
- (iv) There are strong molecular forces of attraction between their molecules.

(b) LIQUID STATE:

- (i) Liquids have no definite shape.
- (ii) They occupies a definite volume of their own.
- (iii) The attractive force between molecules of liquids is lower than solids.
- (iv) They are not absolutely in-compressible, they are compressed to a negligible extent by high pressure.

(c) GASEOUS STATE:

- (i) A gas has no shape of its own.
- (ii) It has no definite volume.
- (iii) It can compress easily on applying pressure.
- (iv) It can diffuse easily.

KINETIC MOLECULAR THEORY OF MATTER:

The Kinetic molecular theory is so named because it deals with that property of the particles which is so crucial in understanding the three states of matter. The important postulates of theory are given below.

Important Postulates:

- (a) Matter is made up of very tiny particles, which are called molecules.
- (b) These molecules are always in motion and possess kinetic energy.
- (c) Their motion can be translational, vibrational or rotational.
- (d) The motion of molecules depends on the state of matter and nature of molecule.
- (e) The molecules of a substance attract each other with a force which depends upon the distance between them.

INTERCONVERSION OF THREE STATES OF MATTER

Most of the important properties of states of matter i.e. solid, liquid and gas depend upon the interconversion of one state into another state.

(a) MELTING POINT:

“The temperature at which solid melts and convert into liquid state is called melting point.”

(b) BOILING POINT:

“The temperature at which liquid converts into gaseous state is called boiling point.”

(c) **FUSION:**

“At melting point the temperature of solid and liquid remains same until all the solid is melted through heat. In this condition the rise in temperature is used in change of state from solid to liquid and this change is called fusion.”

(d) **SUBLIMATION:**

“On heating, some solid substances, which change directly to vapour with out passing through liquid phase. This phenomenon is called sublimation.”

Example of Solid Substances:

- (i) Camphor
- (ii) Iodine
- (iii) Solid carbon dioxide or Dry ice
- (iv) Naphthalene

(e) **EVAPORATION:**

“The process through which change of state from liquid to gas, with out rise in temperature is called evaporation.”

(f) **VAPOURIZATION:**

“The process through which change of state from liquid to gas, by rise in temperature is called vapourization.”

(g) **DIFFUSION:**

“The spreading of substance (gas molecules) through medium like air or liquid is called diffusion.”

GRAHAM’S LAW OF DIFFUSION

Introduction:

In 1833, a Scottish chemist, Thomas Graham studied the rate of diffusion of different gases and formulated Graham’s law.

Statement:

“Rate of diffusion of a gas is inversely proportional to the square root of the density or molar mass of that gas provided the pressure and temperature are the same for the two gases.”

Mathematics Representation:

$$r \propto \frac{1}{\sqrt{d}}$$

in term of molar mass.

$$r \propto \frac{1}{\sqrt{m}}$$

Where,

r = rate (time) of diffusion

d = density of the gas

m = molar mass of the gas

Explanation:

In terms of kinetic molecular theory Graham’s law can be explained as follows:

“Gas molecules vibrates continuously and moving randomly. Because of this, continuous motion molecules passes kinetic energy and applied pressure to the walls of the vessel. As they get space, they escape and transform from one medium to another medium.”

DIFFUSION OF GASES:

In gaseous state, lighter gases can diffuse faster than heavier gases.

DIFFUSION OF LIQUIDS:

Liquid is intermediate between gaseous and solid states like gases liquid molecules are able to move and thus flow and diffuse. The rate of movement of liquid molecules is smaller than gases, hence they diffuse slower than gases.

BROWNIAN MOVEMENT

Introduction:

This property first of all observed by British Botanist, Robert Brown in 1827.

Statement:

“A continuous, rapid, zig-zag motion of suspended particles through medium is called Brownian motion.”

Example:

Consider the following example to understand to Brownian motion clearly.

“Mix some powdered sulphur in water and stir it, after stirring filter the suspended sulphur. Some of the sulphur particles are very small and they can pass through the force of filter paper into filtrate. Now put a drop of this filtrate on a slide and examine it under high powered microscope. It is observed that sulphur particles perform rapid random zig-zag motion through the medium and this motion is called Brownian motion.”

CHAPTER # 07

SOLUTION AND
SUSPENSION

SOLUTION:

“A uniform or homogenous mixture of two or more substance is called a solution.”

COMPONENTS OF SOLUTION

A solution has two components. i.e.

- (a) Solute
- (b) Solvent

(a) **SOLUTE:**

“A solute is a substance which is dissolved and is in smaller quantity.”

(b) **SOLVENT:**

“The substance in which a solute is dissolved and is in larger quantity is called solvent.”

TYPES OF SOLUTION

There are three basic types of solution. i.e.

- (a) Saturated solution
- (b) Un-saturated solution
- (c) Super saturated solution

(a) **SATURATED SOLUTION:**

“A solution which contains maximum amount of solute that it can dissolve at a particular temperature is called saturated solution.”

(b) **UN-SATURATED SOLUTION:**

“A solution which contain less amount of the solute than is required to saturate it at that temperature is called an un-saturated solution.”

(c) **SUPER SATURATED SOLUTION:**

“A solution that is more concentrated than a saturated solution is called super saturated solution.”

TYPES OF SOLUTION ON THE BASIS OF STATES OF MATTER

Since matter exists in three states. i.e.

- (a) Solid
- (b) Liquid
- (c) Gas

On mixing, they produce nine types of solutions, which are listed as following:

TABLE

S.No.	Solute	Solvent	Examples
1.	Gas	Gas	Air (Mixture of 78% N ₂ , 21% O ₂ , and 1% other gases)
2.	Gas	Liquid	Carbonated soft drinks such as coca-cola, Bubble-up etc. NH ₃ gas in water, and air dissolved in water.
3.	Gas	Solid	H ₂ gas adsorbed over palladium metal
4.	Liquid	Gas	Cloud (water vapours in air), steam.
5.	Liquid	Liquid	Alcohol in water, water in milk, milk in tea, vinegar (acetic acid in water)
6.	Liquid	Solid	Amalgam (e.g. Mercury in sodium), and water in jelly powder.
7.	Solid	Gas	Smoke (carbon particles in air),
8.	Solid	Liquid	Sugar in water, sea water
9.	Solid	Solid	Alloys such as brass (copper and zinc) bronze (copper and tin), steel (carbon and iron), glass.

CRYSTALLIZATION:

“The process in which dissolved solute comes out of solution and forms crystals is called crystallization.”

SOLUBILITY:

“The number of grams of the solute necessary to saturate 100 grams of the solvent at that temperature is the solubility of a solute in a solvent at a particular temperature.”

FACTORS AFFECTING SOLUBILITY

The factors that affect the solubility of a solute in a solvent are as follows:

- (a) Temperature
- (b) Pressure
- (c) Nature of solute and solvent

(a) TEMPERATURE:

- (i) Generally it has been observed that the solubility of many solutes in solution increases with the increases in temperature.

Example:

Solubility of Potassium nitrate increases with temperature.

Solubility of sugar (C₆H₁₂O₆) in water,

- at 0°C is 179g/100 ml
- at 100°C is 487g/100 ml

- (ii) The solubility of some solutes also decreases with the increase in temperature.

Example:

- Calcium oxide (CaO) is less soluble in hot water than cold water.

- (iii) Some gases are also more soluble in cold solvent than in hot solvent.

(b) PRESSURE:

- (i) In case of the solutions of solid or liquid solvents, pressure has little effect on their solubility.
- (ii) However, solubility of gases increases with the increase in pressure.

Example:

- Carbon dioxide gas (CO₂) is filled in soda water bottle and in cold drinks (7up, Pepsi etc) under high pressure.

HENRY'S LAW

Statement:

“The solubility of a gas in a liquid is directly proportional to the pressure of gas.”

Mathematically:

$$m \propto P$$

$$m = KP$$

Where,

m = amount of gas dissolved

P = Pressure applied

K = Constant value

(c) NATURE OF SOLUTE AND SOLVENT:

Solute and solvent may be polar (as Alcohol, H₂O) and non-polar (as benzene, petrol).

- (i) The solubility of polar compounds and ionic compounds is more in polar solvents.
- (ii) The solubility of non-polar solutes is more in non-polar solvents.

CONCENTRATION OF SOLUTION

“The amount of a solute dissolved in a specific amount of solvent or solution is called its concentration or strength.”

CONCENTRATED SOLUTION

“A solution containing high concentration of the solute is called concentrated solution.”

DILUTE SOLUTION:

“A solution containing low concentration of the solute is called a dilute solution.”

WAYS OF EXPRESSING THE STRENGTH OF SOLUTION

The following are the common methods of expressing the strength (concentration) of a solution.

i.e.

- (a) Molarity (M)
- (b) Molality (m)
- (c) Mole fraction (x)
- (d) Percentage (%)
- (e) Normality (N). it will not discuss now.

(a) **MOLARITY:**

“The number of moles of solute dissolved in one litre (1 dm³) of a solution (not solvent) is called its Molarity.”

Representative Symbol:

Molarity is denoted by “M”.

Formulae of Molarity:

The formulae by which we can calculate Molarity of the solution are as follows:

(i) **When volume in milli-litres (ml) or cm³**

$$\text{Molarity (M)} = \frac{\text{Mass of solute in gram} \times 1000 \text{ cm}^3 \text{ (ml)}}{\text{Gram formula mass of solution} \times \text{ml (cm}^3\text{) of solution}}$$

(b) **MOLALITY:**

“The number of moles of solute dissolved per 1000g (1 kg) of solvent is called molality.”

Representative Symbol:

Molality is denoted by “m”.

Formulae of Molality:

The formulae by which we can calculate Molality of the solution are as follows:

(i) **When moles is given:**

$$\text{Molality (m)} = \frac{\text{No. of moles of solute}}{\text{Kg. of the solvent}}$$

(ii) **When mass in grams is given:**

$$\text{Molality (m)} = \frac{\text{Mass of solute}}{\text{Gram formula mass}} \times \frac{1000 \text{ gram}}{\text{Gram of solvent}}$$

(c) **MOLE FRACTION:**

“The ratio of moles of solute or solvent to the total moles of the solution is called mole fraction.”

Representative Symbol:

Mole fraction is denoted by “x”.

Formulae of Mole Fraction:

The formulae by which mole fraction can be calculated is,

$$\text{Mole fraction (x)} = \frac{\text{No. of moles of solute}}{\text{Total moles of solute \& solvent in sol}}$$

Important Point of Remember:

- (i) The mole fraction is dimensionless quantity.
- (ii) The sum of mole fractions of all components of a solution must equal to 1.

(d) PERCENTAGE:

This is based on mass (m) or value (v) of components of solution. It is of four types.

(i) Percentage in $\frac{M}{M}$ $\left[\frac{\text{mass}}{\text{mass}} \% \right]$

(ii) Percentage in $\frac{M}{V}$ $\left[\frac{\text{mass}}{\text{volume}} \% \right]$

(iii) Percentage in $\frac{V}{M}$ $\left[\frac{\text{volume}}{\text{mass}} \% \right]$

(iv) Percentage in $\frac{V}{V}$ $\left[\frac{\text{volume}}{\text{volume}} \% \right]$

Examples:

- (i) 5 % (M/M) solution means solute 5 gm in 95 g solvent.
- (ii) 10 % (M/V) solution means solute 10 gm in 100 cm³ solvent.
- (iii) 5 % (V/M) solution means solute 15 cm³ in 100 g solution.
- (iv) 15 % (V/V) solution means solute 15 cm³ in 85 cm³ solvent.

SUSPENSION

“A heterogeneous mixture consists of visible particles, each of which contains many Thousands or even millions of molecules surrounded by molecules of liquid is called suspensions.”

Examples of Suspension:

In our daily life the most common examples of suspension are as follows:

(a) SMOLE:

A suspension of particles of carbon in a gas or air.

(b) MUD (Slime):

A suspension of fine particles of solid in small quantity of liquid.

(c) FOAM (FROTH):

A suspension of fine particles of a gas in a liquid.

(d) EMULSION:

A suspension of droplets of one liquid into another in which it is not soluble.

DIFFERENCE BETWEEN

SOLUTION	SUSPENSION
(a) It is transparent.	It is not transparent.
(b) Particles do not settle down.	Particles settle down.
(c) Components cannot be separated by filtration.	Components can be separated by filtration.
(d) Particles cannot be seen with low power microscope.	Particles can be seen by low power microscope.
(e) It is homogenous.	It is heterogeneous.
(f) The size of particles is between 0.1 to 1nm.	The size of particles is larger than 1000nm.

CHAPTER # 08

ELECTROCHEMISTRY

ELECTROCHEMISTRY:

“It is a branch of chemistry that deals with the relationship between electricity and chemical reactions i.e. with the conversion of electrical energy into chemical energy and vice versa.”

ELECTROLYTES:

“The compounds which ionize or dissociate into negative and positive ions in aqueous solutions and conduct electric current are called electrolytes.”

Example:

- Hydro-chloric acid (HCl)
- Copper sulphate solution (CuSO₄)
- Sodium hydroxide (NaOH)

CLASSIFICATION OF ELECTROLYTES

Electrolytes can be divided into two classes.

- (a) Strong electrolytes
- (b) Weak electrolytes

(a) STRONG ELECTROLYTES:

“The compounds which ionize to a large extent in dilute aqueous solutions and therefore conduct electricity to large extent are called strong electrolytes.”

Examples:

- Solution of Sodium chloride (NaCl)
- Sulphuric acid (H₂SO₄)
- Sodium hydroxide solution (NaOH)

(b) WEAK ELECTROLYTES:

“The compounds that conduct electric current poorly due to ionization to a small extent in aqueous solutions are called weak electrolytes.”

Examples:

- Acetic acid (CH₃COOH)
- Carbonic acid (H₂CO₃)
- Benzoic acid

NON-ELECTROLYTES:

“The compounds which do not ionize in aqueous solutions and thus do not conduct electric current are called non-electrolytes.”

Examples:

- Sugar solution (C₁₂H₂₂O₁₁)
- Glucose (C₆H₁₂O₆)
- Urea (NH₂-CO-NH₂)

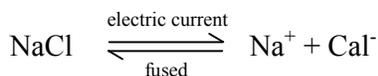
ELECTROLYSIS:

“The chemical change caused by the passage of an electric current through a solution of an electrolyte is called electrolysis.”

EXAMPLES OF ELECTROLYSIS

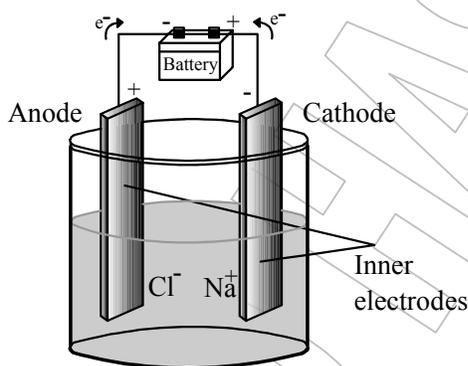
(a) ELECTROLYSIS OF MOLTEN SODIUM CHLORIDE:

Sodium chloride (NaCl), does not conduct electricity in the solid state. To make it conduct electricity, either fuse or melt the salt or dissolve it in water. In the electrolysis of molten NaCl, sodium (Na) moves towards cathode and chlorine gas is given off at the anode, where carbon's (graphite) electrodes are used.



Detail of the Process:

- First prepared the aqueous solution of NaCl (brine) which contains sodium cation (Na^+) and chlorine (Cl^-).
- When electric current is provided in the aqueous solution, the electrolysis process begins.

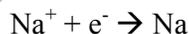


- The sodium ions being positively charged move towards the cathode.
- The chlorine ions (Cl^-) being negatively charged move towards the anode and are discharge to become bubbles of chlorine gas.

Reactions:

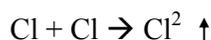
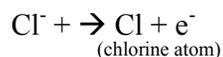
◦ At Cathode:

Sodium ions (Na^+) which are positively charged move towards cathode and gain electrons to get neutralized. Thus Na^+ ions are discharged at cathode and form neutral molten sodium metal.

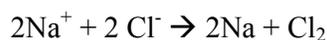


◦ At Anode:

Chloride ions (Cl^-) which are negatively charged move towards anode and lose electrons forming neutral chlorine (Cl) atoms. They rapidly combine to form Cl_2 molecules.

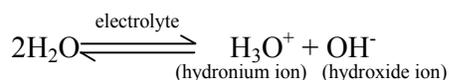


◦ Over-all Reaction:



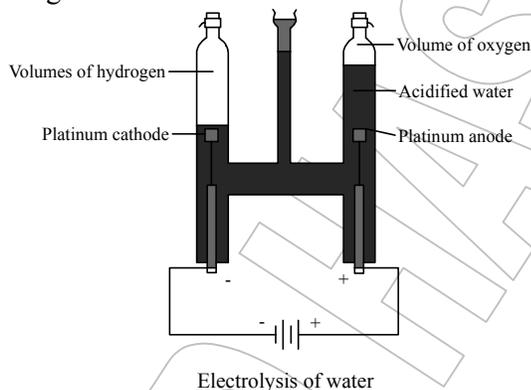
(b) ELECTROLYSIS OF WATER:

Pure water on its own hardly conducts electricity because it ionizes feebly into ions. But on adding of few drops of acid or base or very small quantity of salt is added in water, then it conducts electricity.



Detail of the Process:

- (i) In order to allow the passage of electric current through water, first a small amount of an electrolyte (an acid, a base or a salt) is added.
- (ii) When current is provided in the water containing a little quantity of an electrolyte, the process of electrolysis begins.

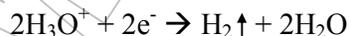


- (iii) The hydronium ions (H_3O^+) being positively charged move towards the cathode and bubble of hydrogen gas is evolved.
- (iv) The hydroxide ions (OH^-) being negatively charged move towards the anode losing the electrons and evolved oxygen gas.

Reactions:

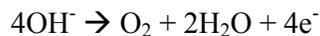
◦ **At Cathode:**

Hydronium ions (H_3O^+) which are positively charged move towards cathode and gain electrons to get neutralized. Thus H_3O^+ ions are discharged at cathode and liberate hydrogen gas.

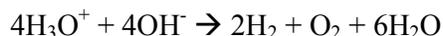


◦ **At Anode:**

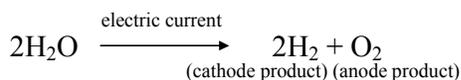
Hydroxide ions (OH^-) which are negatively charged move towards anode and lose electrons get neutralized to liberate oxygen gas.



◦ **Over-all Reaction:**



For simplicity of the electrolysis of water, we write the equation as,



FARADAY'S LAW OF ELECTROLYSIS

Introduction:

Michal Faraday, a well know scientist, in 1833, discovered the quantitative laws governing the process of electrolysis, which are known as Faraday's law of electrolysis.

There are two laws of electrolysis put forward by Faraday.

FARADAY'S FIRST LAW OF ELECTROLYSIS

Statement:

"The amount of any substance deposited or liberated at an electrode during electrolysis is directly proportional to the quantity of current passed through the electrolyte."

Explanation:

If "w" is the weight or amount of a substance deposited or liberated and "A" is ampere of current is passed for "t" seconds, then according to the law,

$$w \propto A \times t$$

or

$$w = ZAt$$

This is mathematical relation of Faraday's first law. Where "Z" is a constant, known as "electro chemical equivalent" of the substance.

ELECTRO-CHEMICAL EQUIVALENT

Statement:

"The weight (amount) of the substance deposited or liberated, when one coulomb of electric charge is passed through an electrolyte is called electro-chemical equivalent of a substance."

Symbolic Representation:

It is denoted by "Z".

Unit:

In S.I. unit, it is expressed in Kg/C. (Kilogram / Coulomb).

Important Point:

- Each element has its own electrochemical equivalent value.

FARADAY'S SECOND LAW OF ELECTROLYSIS

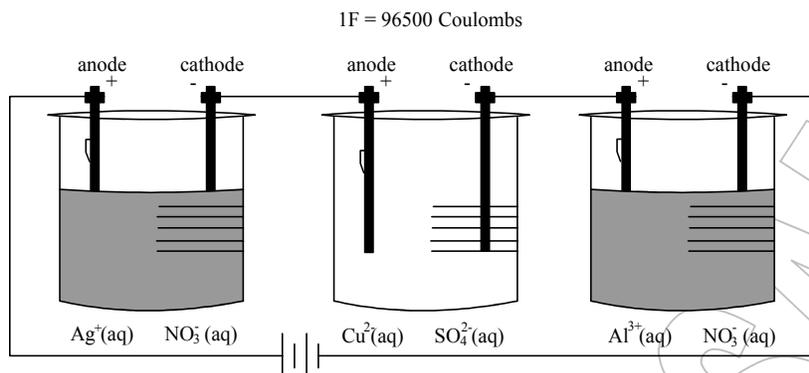
Statement:

"The masses of different substances deposited or liberated, when same quantity of current is passed through different electrolytes, connected in series are proportional to their chemical equivalent masses."

Explanation:

Consider three different electrolytes, i.e. AgNO_3 , CuSO_4 and $\text{Al}(\text{NO}_3)_3$ solution, connected in series. Same quantity of current is passed through them, then the masses of Ag, Cu and Al, deposited on their respective electrodes would be directly proportional to their equivalent masses.

$$\text{Equivalent mass of an element} = \frac{\text{Atomic mass of element}}{\text{Valency of the element}}$$



Electrolytic cells arranged in series

The charge of 96,500 coulombs is called as one Faraday (F) charge.

RELATION BETWEEN EQUIVALENT MASS AND ELECTROCHEMICAL EQUIVALENT

Since 96,500 C (1F) electric charge is required to liberate one gram equivalent mass of a substance, so it is clear that the gram equivalent mass of a substance is 96,500 times greater than its electrochemical equivalent.

$$\text{Gram Equivalent mass} = 96,500 \times \text{E.C.E (Z)}$$

If “e” is the gram equivalent mass and “Z” is the electrochemical equivalent (E.C.E), then we can write it as:

$$E = 96,500 \times Z$$

or

$$e = F \times Z$$

FARADAY:

“The quantity of charge which deposits or liberates exactly one gram equivalent of a substance.”

AMPERE:

“If one coulomb of charge is flowing through a cross-section of a conductor in one second then current is equals to one ampere.”

“It is basic unit of current in SI system.”

$$\text{Ampere} = \frac{\text{Coulomb}}{\text{Time}}$$

COULOMB:

“The quantity of charge when one ampere of current is passed for one second through the circuit.”
It is the SI unit of electric charge.

$$\text{Coulomb} = \text{Ampere} \times \text{Time}$$

or

$$\text{Ampere} = \frac{\text{Coulomb}}{\text{Time}}$$

USES OF ELECTROLYSIS:

Electrolysis is an important process. Some of important uses of electrolysis are listed below:

- (a) This process is used for the extraction of certain metals from their ores.
e.g. Na, Al, Zn, Cu, Mg, Fe etc.
- (b) It is used in electroplating of one metal on other metal.
- (c) It is also used for purification of metals.
e.g. Cu, Al, Agreements, Fe etc.

ELECTROPLATING:

“Electroplating is the process of electrolysis which is used to coat one metal on to another.”

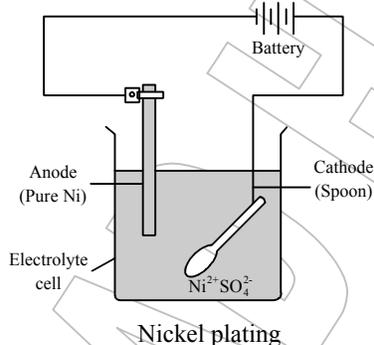
Importance of Electroplating:

- The object to be electroplated is made up of cheaper or bases metal, such as iron, steel etc.
- The cost of the finished product is far less than the objects entirely made of these metals. E.g. Gold coated object is much cheaper than the gold object.
- By this process we protect the metals from rust.

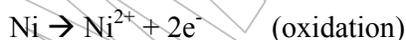
EXAMPLES OF ELECTROPLATING

(a) NICKEL PLATING:

In the process of nickel plating, a piece of pure nickel is used as anode and any substance to be nickel plated is cathode. A solution of nickel sulphate (NiSO_4) is used as the electrolyte in the electrolytic cell. The two electrodes are joined with a battery. On passing the electric current, the anode which is Ni, dissolves in the electrolytic solution forming Ni^{2+} ions by the loss of electrons. Ni^{2+} ions from the solution move towards the cathode, where they gain electrons and are reduced to Ni metal on the surface of substance which is used as cathode.



◦ Reaction at Anode:

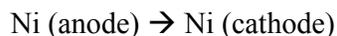


◦ Reaction at Cathode:



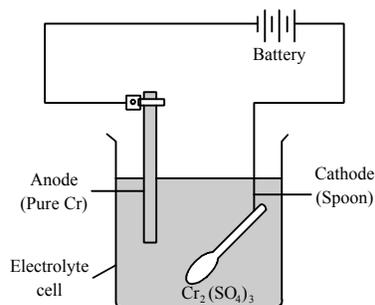
◦ Overall Reaction:

The net reaction is simply the transfer of Ni as Ni^{2+} through NiSO_4 solution towards the cathode and get it coated with Ni metal on the surface. The sum of reduction and oxidation is:



(b) CHROMIUM PLATING:

In the process of chromium plating, a piece of pure chromium is used as anode and any substance to be chromium plated is cathode. A solution of chromium sulphate ($\text{Cr}_2(\text{SO}_4)_3$) is used as the electrolyte in the electrolytic cell. The two electrodes are joined with a battery. On passing the electric current, the anode which is Cr, dissolves in the electrolytic solution forming Cr^{3+} ions by the loss of electrons. Cr^{3+} ions from the solution move towards the cathode, where they gain electrons and are reduced to Cr metal on the surface of substance which is used as cathode.



◦ **Reaction at Anode:**

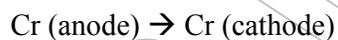


◦ **Reaction at Cathode:**



◦ **Overall Reaction:**

The net process is simply to transfer of Cr as Cr^{3+} ions through the solution $\text{Cr}_2(\text{SO}_4)_3$, towards cathode and coated it with Cr metal. The sum of oxidation and reduction is:



ELECTROCHEMICAL CELLS:

“The cell which is used to convert chemical energy into electrical energy or vice versa is called electrochemical cells.”

VOLTAIC CELL:

“An electrochemical cell which converts chemical energy into electrical energy is known as voltaic cell.”

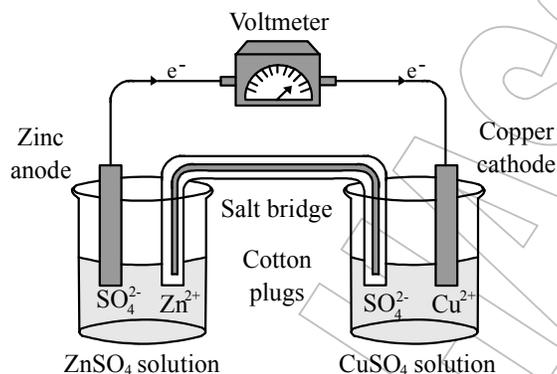
Voltaic cell also called “Galvanic cell”.

DANIELL CELL

“A Daniell cell is the simplest of the Galvanic or Voltaic cells which is used to convert chemical energy into electrical energy spontaneously.”

Construction:

Daniell cell consists of two half cells. One half cell is copper (Cu-metal) dipped in 1M CuSO_4 solution and the other half cell is Zinc rod (Zn-metal) dipped in 1M ZnSO_4 solution. The two half cells should be separated from each other by a porous partition or a salt bridge. The two half cells or single electrodes are connected together to form a complete cell.



Daniell cell (A Zn - Cu voltaic cell)

Working:

When both cell with electrodes are connected externally through a voltmeter by means of metal wire. The cell starts producing electric current at once. Zn undergoes oxidation to form Zn^{2+} ions by the loss of 2-electrons to go into ZnSO_4 solution Zn acts as anode. The electrons which are free at Zn-electrode travel through the wire externally to Cu-electrodes. These electrons are accepted by Cu^{2+} ions of CuSO_4 solution and Cu^{2+} ions undergo reduction to deposit copper metal at Cu-electrode which acts as cathode.

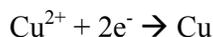
In the whole process Zn-electrode dissolves in the solution of ZnSO_4 and reduces in size, while copper electrode grows in size due to the deposition of Cu-metal and cell works.

Cell Reactions:

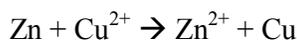
◦ At Anode:



◦ At Cathode:



◦ Overall Reaction:

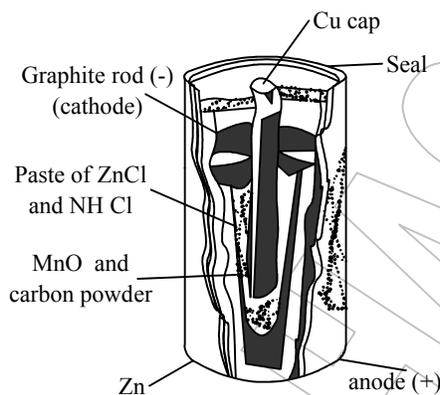


DRY CELL

“A primary cell, which is used to convert chemical energy into electrical energy is called a dry cell.”

Construction:

It consists of a zinc container act as negative electrode filled with a paste of ammonium chloride (NH_4Cl), manganese dioxide (MnO_2), powder carbon and water. A carbon rod is placed in the centre and acts as positive electrode. The whole cell is covered with a safety cover.



Dry cell

Working:

When zinc and graphite electrodes are connected by a metallic wire, Zn gets oxidized to form Zn^{2+} ions which pass into the wet paste leaving behind electrons on the Zn container and the electrons move from Zn electrode to carbon electrode through the external circuit.

The cell reactions are complex.

Uses of Dry Cell:

Dry cells are used in torches, radio sets, clocks, flash lights, calculators, transistors appliances etc.

KEY POINT:

Primary Cells:

“Cell whose working does not depends upon any external source.”

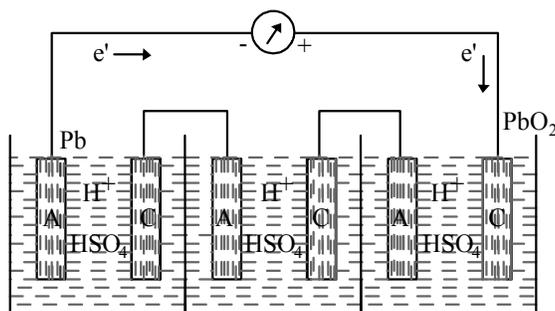
LEAD-STORAGE BATTERY

“It is a secondary battery and is a reversible cell which can be restored to its original condition. It is also called motor-battery.”

Construction:

In lead-storage battery, there are several anode and several cathodes which are connected together in series; about six cells are connected together. Each cell has a voltage of 2V.

In lead-storage battery anodes are the lead alloy and cathodes are made up of red lead oxide (PbO_2). The electrolyte is dilute Sulphuric acid, which having concentration of 30%.



Lead storage battery

Working:

As the cell reaction proceeds PbSO₄ (s) precipitates and partially coats both the electrodes, the water formed dilutes the H₂SO₄. The battery is said to be discharged. Now by connecting the battery to an external electrical source, we can force the electrons to flow in the opposite direction, i.e. the net cell reaction can be reversed and the battery is recharged.

Uses of Lead-Storage Battery:

Lead storage battery is used in automobiles. It is also used to perform many electrolytical processes.

MUHAMMAD HASSAM

CHAPTER # 09

ACIDS, BASE AND SALTS

ACIDS:

“The compound which provided hydrogen ion (H^+) in aqueous solutions are called acids.”

OR

“A compound containing hydrogen which may be replaced directly or indirectly by a metal is called acid.”

CLASSES OF ACIDS:

Acids can be classified on the basis of their sources. There are two main classes of acids.

- (1) Organic acids
- (2) Mineral or Inorganic acids.

(1) ORGANIC ACIDS:

Organic acids occur as natural products in organic matter (i.e. plant or animal materials).

Examples:

ORGANIC ACIDS	SOURCES
Acetic acid (ethanoic acid)	Vinegar
Lactic acid	Milk
Citric acid	Lime, lemon
Tartaric acid	Grapes
Amino acid	Proteins
Fatty acid	Fats and oils
Ascorbic acid (Vitamin C)	oranges

(2) MINERAL OR INORGANIC ACIDS:

These acids can be prepared in the laboratory from mineral elements or inorganic matter.

Examples:

ACIDS	FORMULA	CONSTITUENT ELEMENTS
Hydrochloric acid	HCL	Hydrogen, Chlorine
Sulphuric acid	H_2SO_4	Hydrogen, Sulphur, Oxygen
Nitric acid	HNO_3	Hydrogen, Nitrogen, Oxygen
Carbonic acid	H_2CO_3	Hydrogen, Carbon, Oxygen
Phosphoric acid	H_3PO_4	Hydrogen, Phosphorus, Oxygen

PHYSICAL STATES OF ACIDS:

Generally the acids exist in liquid state but a few are also found in gaseous and solid state.

Example:

- SOLID STATE:** Lactic acid, Oxalic acid ($H_2C_2O_4$)
- LIQUID STATE:** Sulphuric acid, Acetic acid (CH_3COOH)
- GASEOUS STATE:** Hydrobromic acid (HBr)

CONCEPTS OF ACIDS

A brief review of the various concepts regarding acids are given.

THE ARRHENIUS CONCEPT

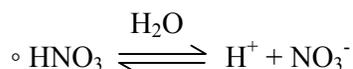
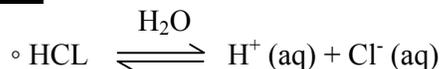
Introduction:

In 1887, a Swedish scientist "S.A. Arrhenius" purposed the following concept regarding acid.

Statement:

"An acid is a substance which produces H^+ ions (hydrogen ions) in aqueous solution."

Example:



Conclusion:

From the above examples it is clear that the Arrhenius concept is based upon ionic dissociation of acids in water.

KEY POINT:

Ionization:

"The process of formation of H^+ ion from acid in the aqueous solution is called ionization."

LOWRY AND BRONSTED CONCEPT

Introduction:

In 1923, two scientists Lowry and Bronsted gave a new concept of acids.

Statement:

"Acid is a species (a compound or ion) which donates or tends to donate a proton (H^+)."

Example:



Conclusion:

From the above two examples, it is obvious that according to Lowry and Bronsted concept, proton donar (acid) and proton-accepter (base) must exit. But in those reactions where this condition doesn't exist this concept can not be applied.

Example:

- $\text{Zn (s)} + 2\text{HCl (aq)} \rightarrow \text{ZnCl}_2 \text{ (aq)} + \text{H}_2 \text{ (g)}$
- $2\text{Al (s)} + 6\text{HCl (aq)} \rightarrow 2\text{AlCl}_3 \text{ (aq)} + 3\text{H}_2 \text{ (g)}$
- $\text{Mg (s)} + 2\text{HNO}_3 \text{ (aq)} \rightarrow \text{Mg(NO}_3)_2 \text{ (aq)} + \text{H}_2 \text{ (g)}$

(c) Reaction with Carbonates:

Most of acids decompose carbonates with evolution of CO_2 gas and water is formed.

Example:

- $2\text{HCl (aq)} + \text{Na}_2\text{CO}_3 \text{ (aq)} \rightarrow 2\text{NaCl (aq)} + \text{H}_2\text{O (l)} + \text{CO}_2 \text{ (g)}$
- $2\text{HNO}_3 \text{ (aq)} + \text{ZnCO}_3 \text{ (aq)} \rightarrow \text{Zn(NO}_3)_2 \text{ (aq)} + \text{H}_2\text{O (l)} + \text{CO}_2 \text{ (g)}$
- $\text{H}_2\text{SO}_4 \text{ (aq)} + \text{CaCO}_3 \text{ (aq)} \rightarrow \text{CaSO}_4 \text{ (aq)} + \text{H}_2\text{O (l)} + \text{CO}_2 \text{ (g)}$

(d) Reaction with Bi-Carbonates:

Most of the acids decompose bi-carbonates with evolution of CO_2 gas and H_2O is formed.

Example:

- $\text{HCl} + \text{NaHCO}_3 \rightarrow \text{NaCl} + \text{H}_2\text{O} + \text{CO}_2$

(e) Reaction with Metal Oxide:

Acids react with metal oxides in the same way as with metal hydroxide. They produce salt and water. But oxides are usually more reactive.

Example:

- $6\text{HCl} + \text{Fe}_2\text{O}_3 \rightarrow 2\text{FeCl}_3 + 3\text{H}_2\text{O}$
- $3\text{HCl} + \text{Fe(OH)}_3 \rightarrow \text{FeCl}_3 + 3\text{H}_2\text{O}$
(Iron hydroxide)

BASES:

“A base is a substance which provides hydroxyl ions (OH^-) in aqueous solution.”

OR

“A base is any substance that can accept or combine with positive hydrogen ions (or proton).”

CLASSES OF BASES

Bases can be classified on the basis of their solubility in water. There are two main classes of bases.

- (1) Alkalis (2) Normal Bases.

(1) ALKALIS:

“An alkali is a basic hydroxide which is soluble in water.”

Example:

- Sodium hydroxide (NaOH)
- Potassium hydroxide (KOH)

(2) NORMAL BASES:

“Normal bases are those hydroxides which are partially soluble or insoluble in water.”

Example:

- Calcium hydroxide [Ca(OH)_2]
- Aluminium hydroxide [Al(OH)_3]
- Copper hydroxide [Cu(OH)_2]
- Iron hydroxide [Fe(OH)_3]

KEY POINT:

Metal Hydroxide:

These are bases which are prepared by heavy metals. This is insoluble in H₂O.

e.g. Fe(OH)₃, Cu(OH)₂, Ni(OH)₂, Co(OH)₂.

PHYSICAL STATE OF BASES:

Generally the bases exist in liquid state but a few are also found in solid state.

Example:

(a) **SOLID STATE:** Sodium Hydroxide (NaOH) as caustic.

(b) **LIQUID STATE:** Calcium Hydroxide, Potassium Hydroxide.

CONCEPT OF BASES

A brief review of the various concepts regarding bases are given.

THE ARRHENIUS CONCEPT

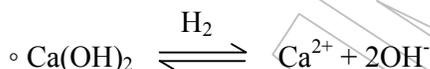
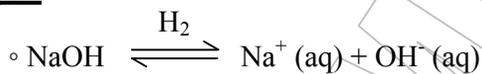
Introduction:

In 1887, a Swedish scientist "S.A. Arrhenius" proposed the following concept regarding base.

Statement:

"An base is a substance which gives OH⁻ ions (hydroxyl ions) in aqueous solution."

Example:



Conclusion:

From the above examples it is clear that the Arrhenius concept is based upon ionic dissociation of bases in water.

LOWRY AND BRONSTED CONCEPT

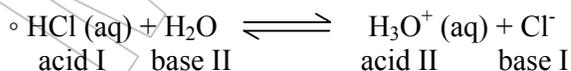
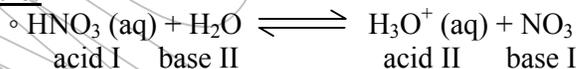
Introduction:

In 1923, two scientists Lowry and Bronsted gave a new concept of bases.

Statement:

"A base is a species (a compound or ion) which accepts or tends to accept a proton (H⁺)."

Example:



Conclusion:

From the above two examples, it is obvious that according to Lowry and Bronsted concept, proton donor (acid) and proton acceptor (base) must exist. But in those reactions where this condition doesn't exist this concept can not be applied.

LEWIS CONCEPT

Introduction:

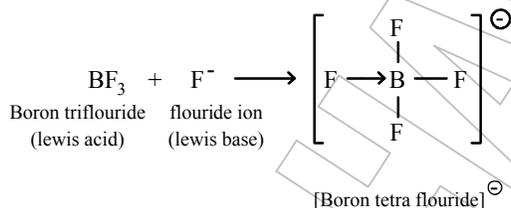
In 1923, an American chemist, Lewis put forward a more general concept of bases.

Statement:

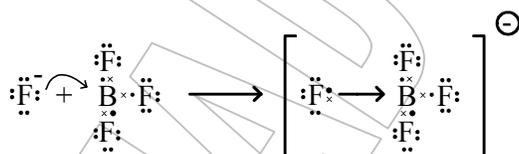
“A base is a species (molecule or ion) which can donate a pair of electrons.”

A base is also called an “nucleophile” (nucleus loving).

Example:



Dot & Cross Structure: (Lewis Structure)



Conclusion:

From the above example it is clear that shared pair of electrons is provided by F⁻ (Lewis base) to BF₃ (Lewis acid) and co-ordinate covalent bond is formed.

PROPERTIES OF ACIDS

Bases are compounds which exhibit the following properties in aqueous solution.

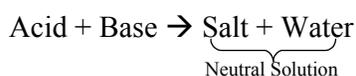
PHYSICAL PROPERTIES:

- Alkalis have a sour taste.
- Bases are soapy to the touch.
- They turn red litmus to blue, methyl orange to yellow and colourless phenolphthalein in pink.
- They conduct electricity.
- Concentrated forms of the two caustic alkalis, NaOH and KOH, are corrosive.

CHEMICAL PROPERTIES:

(a) Reaction with Acids:

All bases react with acid to form salts and water only. This type of reaction called “neutralization.”



Example:

- $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
- $2\text{HCl} + \text{Ca}(\text{OH})_2 \rightarrow \text{CaCl}_2 + \text{H}_2\text{O}$
- $\text{H}_2\text{SO}_4 + 2\text{KOH} \rightarrow \text{K}_2\text{SO}_4 + \text{H}_2\text{O}$

(b) Reaction with Metals:

Bases dissolve certain metals like tin, zinc, aluminium etc.

Example:

- $\text{Zn} (\text{s}) + \text{NaOH} (\text{aq}) + \text{H}_2\text{O} (\text{l}) \rightarrow \text{Na}_2\text{ZnO}_2 (\text{aq}) + 2\text{H}_2 (\text{g})$
(sodium aluminate)
- $2\text{Al} (\text{s}) + 2\text{NaOH} (\text{aq}) + 2\text{H}_2\text{O} (\text{l}) \rightarrow 2\text{NaAlO}_2 (\text{aq}) + 3\text{H}_2 (\text{g})$
(sodium zincate)

(c) Reaction with Non-Metals:

Bases also dissolve certain non-metals like silicon, but the reaction is slow at room temperature.

Example:

- $\text{Si} (\text{s}) + 2\text{NaOH} + \text{H}_2\text{O} (\text{aq}) \rightarrow \text{Na}_2\text{SiO}_3 (\text{aq}) + 2\text{H}_2 (\text{g})$
(sodium silicate)

(d) Reaction with Heavy Metals Salts:

Bases precipitate hydroxides of heavy metals from their salt solutions.

Example:

- $\text{FeCl}_3 + 3\text{NaOH} \rightarrow \text{Fe}(\text{OH})_3 + 3\text{NaCl}$
- $\text{CrCl}_3 + \text{KOH} \rightarrow \text{Cr}(\text{OH})_3 + 3\text{KCl}$

(e) Reaction with Ammonium Salts:

Alkalis react with ammonium salt when we give a sufficient amount of heat, give ammonia gas.

Example:

- $\text{NaOH} + \text{NH}_4\text{NO}_3 \rightarrow \text{NaNO}_3 + \text{H}_2\text{O} + \text{NH}_3$
(Ammonium Nitrate) (Ammonia)

(f) Reaction with Organic Compounds:

Bases react with fats to form soaps. They are able to dissolve proteins and certain organic compounds.

ACIDITY OF BASES:

“The number of ionizable hydroxyl ion (OH^-) present in a molecule of a base is called acidity of base.”

The following table gives the Acidity of some bases

Base	Ions Produced	Acidity
NaOH	$\text{Na}^+ + \text{OH}^-$	1 or mono-acid
$\text{Ca}(\text{OH})_2$	$\text{Ca}^{2+} + 2\text{OH}^-$	2 or di-acid
$\text{Fe}(\text{OH})_3$	$\text{Fe}^{3+} + 3(\text{OH})^-$	3 or tri-acid

BASICITY OF ACIDS:

“The number of ionizable hydrogen atoms present in the molecule of an acid is called basicity of acid.”

The following table gives the Basicity of some acids

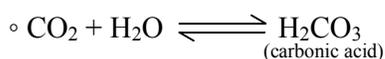
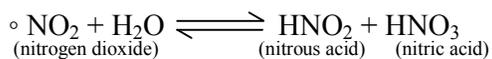
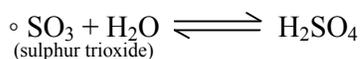
Acid	Ions Produced	Basicity
HCL	$H^+ + Cl^-$	1 or mono-basic
H_2SO_4	$2H^+ + SO_4^{2-}$	2 or di-basic
H_3PO_4	$3H^+ + PO_4^{3-}$	3 or tri-basic

PREPARATIONS

ACIDS FROM NON-METAL OXIDES:

Actually non-metal oxides are acidic in nature and dissolve it in water to produce acids.

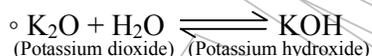
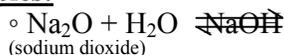
Examples:



BASES FROM METAL OXIDES:

Normally the metal oxides are basic in nature and produce bases with water.

Examples:



AMPHOTERIC OXIDES:

These are the oxides which have both acidic and basic character.

e.g. Al_2O_3 (Aluminium Oxide)

STRENGTH OF ACIDS:

The strength of an acid depends on the concentration of hydrogen ions (H^+) in water.

On the basis of strength acids are classified into two main types.

- Strong acids
- Weak acids

(a) STRONG ACIDS:

“Those acids which are completely ionize in aqueous solution called strong acids.”

Examples: HCL, HNO_3 , H_2SO_4 .

(b) WEAK ACIDS:

“Those acids which ionize to a small extent in water called weak acids.”

Examples: Acetic acid (CH₃COOH), Carbonic acid (H₂CO₃).

Important Note:

- 100% HCL ionizes up to 99.9%.
- 100% CH₃COOH ionizes up to only 1.3%.

STRENGTH OF BASES:

The strength of a base depends on the concentration of hydroxyl ions (OH⁻) in water.

On the bases of strength, bases are also classified into two main types.

- (a) Strong bases
- (b) Weak bases

(c) STRONG BASES:

“Those bases which are completely ionize in aqueous solution, are called strong bases.”

Examples: NaOH, KOH.

(d) WEAK BASES:

“Those bases which ionize into small extent in aqueous solution, called weak bases.”

Examples: Al(OH)₃, NH₄OH.

SALTS:

“A salt is neutral compound formed due to neutralization of an acid by a base.”

HYDROLYSIS OF SALTS:

“Partial decomposition of a salt, due to its reaction with water, into its constituent acid and base or any other substance by any chemical reaction with water is called Hydrolysis of salts.”

OR

“The process in which a salt is decomposed water to produce an acid and a base is called Hydrolysis of Salt.”

TYPES OF SALTS

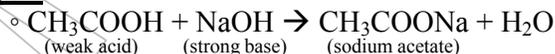
According to the process of hydrolysis, salts are mainly divided into three categories.

- (a) Basic salts
- (b) Acidic salts
- (c) Neutral salts

(a) BASIC SALTS:

“The salts formed by the combination of weak acids and strong base are called basic salts.”

Example:



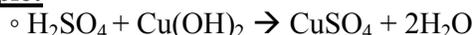
- When basic salts dissolve in water it gives strong base and weak acid.



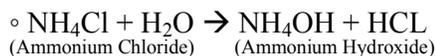
(b) ACIDIC SALTS:

“The salts formed by the combination of strong acids and weak base are called acidic salts.”

Example:



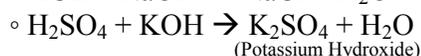
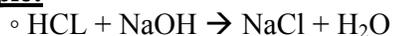
◦ When acidic salts dissolve in water it gives strong acid and weak base.



(c) NEUTRAL SALTS:

“The salts formed by the combination of a strong acids with a strong base called neutral salts.”

Example:



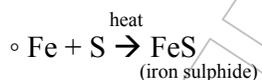
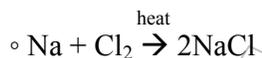
PREPARATION OF SALTS

Salts can be prepared by a large number of methods. Some of the important methods is given below.

(a) BY REACTION OF METALS WITH NON-METALS:

When metals heated up with non-metals it gives respective salts.

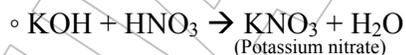
Example:



(b) BY NEUTRALIZATION REACTION:

Acids neutralize bases and formed salts.

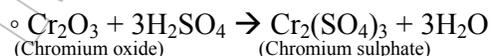
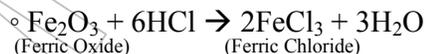
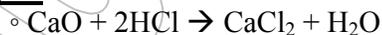
Example:



(c) BY REACTION OF ACIDS WITH METALLIC OXIDES:

When acids react with metallic oxides it formed respective salts.

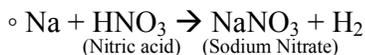
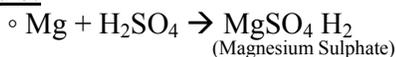
Example:



(d) BY REACTION OF ACIDS WITH METALS:

When acids react with metals they form respective salts and hydrogen gas evolved.

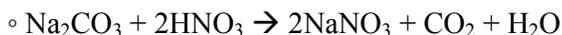
Example:



(e) BY REACTION OF ACIDS WITH METAL CARBONATES:

When acids react with metallic carbonates, they formed respective salts.

Example:



(f) BY REACTION OF ACIDS WITH Bi-CARBONATES:

When acids react with metallic bicarbonates, they formed respective salts.

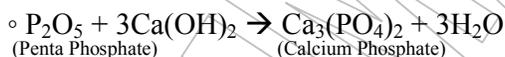
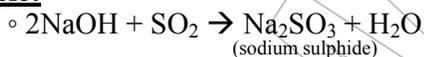
Example:



(g) BY REACTION OF BASES WITH NON-METAL OXIDES:

When strong bases react with non-metallic oxides they formed their respective salts.

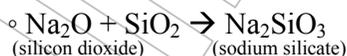
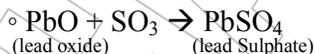
Example:



(h) BY REACTION OF METAL OXIDE WITH NON-METAL OXIDE:

When metallic oxides combine with non-metallic oxides, it formed respective salts.

Example:



pH SCALE

Introduction:

In 1909, a well known chemist Sorenson invented pH meter to measure the concentration of H^+ ion in aqueous solution.

Statement:

“pH is the negative logarithm to the base 10 of the molar concentration of hydrogen ions.”

Formula:

pH of the solution can be find out by using formula.

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

KEY POINT:

pH means power of hydrogen ion (H⁺ ion) concentration.

pOH SCALE**Statement:**

“pOH of a solution is the negative logarithm to the base 10 of the molar concentration of hydroxyl ion (OH⁻ ion).

Formula:

pOH of the solution can be find out by using formula.

$$\text{pOH} = -\log [\text{OH}^-]$$

◦ For various H⁺ ions concentration the values of pH and pOH are given in the following table.

Key Points:

- In pH scale 1 → 6.9 is acidic medium.
- In pH scale at 7.0 is neutral medium.
- In pH scale 7.1 → 14 is basic medium.
- pOH means power of hydroxyl ion (OH⁻ ion) concentration.

[H ⁺]	pH	pOH	Nature of Solution
10 ⁰	0	14	} → Strongly acidic
10 ⁻¹	1	13	
10 ⁻²	2	12	
10 ⁻³	3	11	} → Weakly acidic
10 ⁻⁴	4	10	
10 ⁻⁵	5	9	
10 ⁻⁶	6	8	
10 ⁻⁷	7	7	→ Neutral
10 ⁻⁸	8	6	} → Weakly basic
10 ⁻⁹	9	5	
10 ⁻¹⁰	10	4	
10 ⁻¹¹	11	3	
10 ⁻¹²	12	2	} → Strongly basic
10 ⁻¹³	13	1	
10 ⁻¹⁴	14	0	

CRYSTALLIZATION:

“The process in which crystals of a compound is formed, called process of crystallization.”

MUHAMMAD HASSAM

CHAPTER # 10

CHEMICAL ENERGETICS

THERMO-CHEMISTRY:

“The branch of chemistry which deals with the study of heat changes in chemical reactions is called thermochemistry.”

THERMO-CHEMICAL REACTION:

“The chemical reactions during which material changes are accompanied with change in heat energy are called thermochemical reaction.”

TYPES OF THERMOCHEMICAL REACTION

There are two types of thermochemical reactions.

- (1) Exothermic reactions
- (2) Endothermic reactions

(1) EXOTHERMIC REACTION:

“The reaction which evolves heat instead of giving it in is called exothermic reaction.”

OR

“The reaction in which heat is evolved is called exothermic reactions.”

Explanation:

“Exo” means “outside” and “thermic” means “heat”. Hence “Exothermic reactions” are those reactions in which heat is evolved. In such case potential energy of substance decreases.

Examples:

(a) Combustion of Coal:

The combustion of coal in air is the best example of exothermic reaction. About 393.7 Kilo Joules of heat energy is released, when 1 mole of coal is burnt in 1 mole of oxygen to produce 1 mole of CO₂.



(b) Burning of Methane:

Burning of methane in presence of oxygen is another example of exothermic reaction. When 1 mole of methane is burnt in 2 mole of oxygen then one mole of CO₂ and 2 moles of H₂O are formed. During this reaction 890 kilo joules / mole of heat energy evolved.



(2) ENDOTHERMIC REACTION:

“The reaction which absorbs heat instead of giving it out is called endothermic reaction.”

OR

“The reaction in which heat is absorbed is called endothermic reactions.”

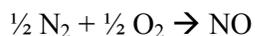
Explanation:

“Endo” means “inside” and “thermic” means “heat”. i.e. in “Endothermic reactions” heat is absorbed inside during chemical reaction. In such case the products have higher potential energy than the reactants.

Examples:

(a) Formation of Nitric Oxide:

1 mole of nitric oxide (NO) is formed by combination of $\frac{1}{2}$ mole of N_2 and $\frac{1}{2}$ mole of O_2 . This is the example of endothermic reaction and heat absorbed is about + 90.25 KJ/mole.



$$\Delta H = +90.25 \text{ KJ/mol.}$$

ENTHALPY

“It is the thermodynamics state function in which energy given out or absorbed at constant pressure.”

Symbolic Representation:

“Enthalpy” is represented by “H”.

ENTHALPY FOR EXOTHERMIC REACTIONS

In exothermic reactions the enthalpy of products is less than the enthalpy of reactants. Its means exothermic reactions ΔH is -ve (negative).

ENTHALPY FOR ENDOTHERMIC REACTIONS

In endothermic reactions the enthalpy of reactants is less than the enthalpy of products. Its means endothermic reactions ΔH is +ve (positive).

◦ $\Delta H = (H_2 - H_1)$

◦ Change of Enthalpy = (heat content of products – heat content of reactants)

HEAT OF NEUTRALIZATION

“The amount of heat released during a neutralization reaction in which 1 mole of water is formed is called as the heat of neutralization.”

OR

“The amount of heat released when 1 mole of hydrogen ions (H^+) from an acid reacts with 1 mole of hydroxyl ions (OH^-) from a base to form salt and one mole of water is called the heat of neutralization.”

Neutralization Reaction:

“The reaction between acid and base to form a salt and water.”