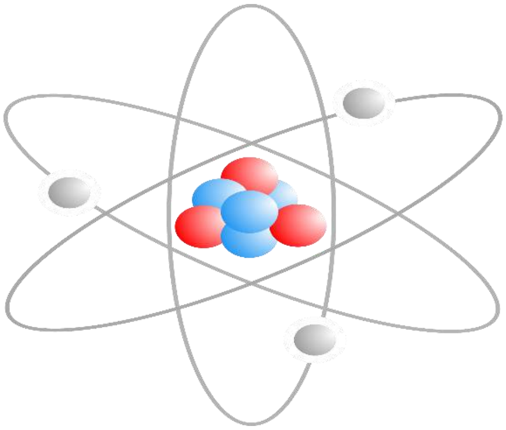
ENGINEERRING CHEMISTRY

(FOR DIPLOMA ENGINEERING)



Prepared BY
Dipti Laxmi Bhuyan
(Sr. Lecturer in Chemistry)
U.C.P Engg. School, Berhampur

SYLLABUS

(As per SCTE&VT)

A. PHYSICAL CHEMISTRY

Chapter 1: Atomic structure : Fundamental particles (electron, proton & neutron Definition, mass and charge).Rutherford's Atomic model (postulates and failure), Atomic mass and mass number, Definition, examples and properties of Isotopes, isobars and isotones. Bohr's Atomic model (Postulates only), Bohr-Bury scheme, Aufbau's principle, Hund's rule, Electronic configuration (up to atomic no 30).

Chapter 2 : Chemical Bonding : Definition , types (Electrovalent, Covalent and Coordinate bond with examples (formation of NaCl, $MgCl_2$, H_2,Cl_2, O_2, N_2, H_2O, CH_4, NH_3, $NH_4{}^+$, SO_2).

Chapter 3 : Acid base theory : Concept of Arrhenius, Lowry Bronsted and Lewis theory for acid and base with examples (Postulates and limitations only). Neutralization of acid & base. Definition of Salt, Types of salts (Normal, acidic, basic, double, complex and mixed salts, definitions with 2 examples from each).

Chapter 4: Solutions : Definitions of atomic weight, molecular weight, Equivalent weight. Determination of equivalent weight of Acid, Base and Salt.
Modes of expression of the concentrations (Molarity , Normality & Molality) with Simple Problems. pH of solution (definition with simple numericals)
Importance of pH in industry (sugar, textile, paper industries only)

Chapter 5 : Electrochemistry : Definition and types (Strong & weak) of Electrolytes with example. Electrolysis (Principle & process) with example of NaCl (fused and aqueous solution). Faraday's 1st and 2nd law of Electrolysis (Statement, mathematical expression and Simple numerical) Industrial application of Electrolysis- Electroplating (Zinc only)

Chapter 6 : Corrosion: Definition of Corrosion, Types of Corrosion- Atmospheric Corrosion, Waterline corrosion. Mechanism of rusting of Iron only. Protection from Corrosion by (i) Alloyingand (ii) Galvanization.

B.INORGANIC CHEMISTRY

Chapter 7 : Metallurgy: Definition of Mineral, ores , gangue with example. Distinction between Ores And Minerals. General methods of extraction of metals,

i) Ore Dressing
ii) Concentration (Gravity separation, magnetic separation, Froth floatation & leaching)
iii) Oxidation (Calcinations, Roasting)
iv) Reduction (Smelting, Definition & examples of flux, slag)
v) Refining of the metal (Electro refining, & Distillation only)

Chapter 8 : Alloys: Definition of alloy. Types of alloys (Ferro, Non Ferro & Amalgam) with example. Composition and uses of Brass, Bronze, Alnico, Duralumin

C. ORGANIC CHEMISTRY

Chapter 9: Hydrocarbons : Saturated and Unsaturated Hydrocarbons (Definition with example)

Aliphatic and Aromatic Hydrocarbons (Huckle's rule only). Difference between Aliphatic and aromatic hydrocarbons
IUPAC system of nomenclature of Alkane, Alkene, Alkyne, alkyl halide and alcohol (up to 6 carbons) with bond line notation.
Uses of some common aromatic compounds (Benzene, Toluene, BHC, Phenol, Naphthalene, Anthracene and Benzoic acid) in daily life.

D. INDUSTRIAL CHEMISTRY

Chapter 10 : **Water Treatment** : Sources of water, Soft water, Hard water, hardness, types of Hardness (temporary or carbonate and permanent or non-carbonate), Removal of hardness by lime soda method (hot lime & cold lime—Principle, process & advantages) , Advantages of Hot lime over cold lime process.
Organic Ion exchange method (principle, process, and regeneration of exhausted resins) **Chapter 11 : Lubricants**: Definition of lubricant, Types (solid, liquid and semisolid with examples only) and specific uses of lubricants (Graphite, Oils, Grease), Purpose of lubrication **Chapter 12 : Fuel**: Definition and classification of fuel, Definition of calorific value of fuel, Choiceof good fuel.
Liquid: Diesel, Petrol, and Kerosene --- Composition and uses.
Gaseous: Producer gas and Water gas (Composition and uses). Elementary idea about LPG, CNG and coal gas (Composition and uses only).

Chapter 13 : Polymer: Definition of Monomer, Polymer, Homo-polymer, Co-polymer and Degree of polymerization. Difference between Thermosetting and Thermoplastic, Compositionand uses of Polythene, & Poly-Vinyl Chloride and Bakelite.
Definition of Elastomer (Rubber). Natural Rubber (it's draw backs). Vulcanisation of Rubber. Advantages of Vulcanised rubber over raw rubber.

Chapter 14: Chemicals in Agriculture: Pesticides: Insecticides, herbicides, fungicides-Examples and uses.

Bio Fertilizers: Definition, examples and uses.

CHAPTER-1

(ATOMIC STRUCTURE)

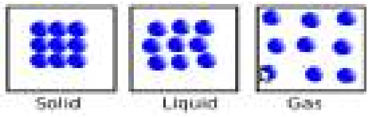
ATOMIC STRUCTURE

LEARNING OBJECTIVES

- FUNDAMENTAL PARTICLES
- RUTHERFORD'S ATOMIC MODEL(POSTULATES & FAILURE)
- ATOMIC MASS AND MASS NUMBER
- ISOTOPE, ISOBAR & ISOTONE
- BOHR'S ATOMIC MODEL
- BOHR-BURY SCHEME
- AUF-BAU PRINCIPLE
- HUND'S RULE
- ELECTRONIC CONFEGURATION

HISTORY:

In 1808 John Dalton said that all matter are made up of small, indivisible particles known as 'atoms'. In nineteenth century it was found that atom cannot be subdivided further. However the experiment conducted by Rutherford, Mosley, Bohr and other revealed that atom is not the smallest particle & it can be subdivided further into three fundamental particles (electron, proton & neutron)



(Fig.1 Arrangement of atoms in three states of matter)

FUNDAMENTAL PARTICLES:

The sub-atomic particles 'electrons, protons and neutrons' are called fundamental particles of all matters.

(i) Electron:

- Discovered by J. J. Thomson in Cathode ray experiment.
- It is designated by the symbol e^- .
- It is a negatively charged particle having value -1.6×10^{-19} coulombs.
- The mass of an electron is approximately 9.11×10^{-31} kilograms.
- Radius of electron is 10^{-15} cm.

(ii) Proton:

- Discovered by E. Goldstein in anode rays experiment.
- It is designated by the symbol ' p' , H^+
- It is a positively charged particle having value $+1.6\times10^{-19}$ coulombs.
- The mass of a proton is approximately 1.672×10^{-27} kilograms.

(iii) Neutron:

- Discovered by James Chadwick. .
- It is designated by the symbol ‘ n’ .
- It is a neutrally charged particle i.e have no charge.
- The mass of a neutron is approximately 1.675×10^{-27} Kg kilograms.

Fundamental Particle	**Mass**	**Charge**	**Relative Charge**
Electron	9.11×10^{-31} kg	-1.602×10^{-19} Coulomb	– 1
Proton	1.672×10^{-27} Kg	$+1.602 \times 10^{-19}$ Coulomb	+1
Neutron	1.675×10^{-27} Kg	0	0

NOTE:

After discovery of electron & proton, J. J. Thomson was the first to give an atomic model called Thomson’s atomic model. According to him in a positively charged sphere, negatively charged electrons are embedded in it. This model is also known as "Plum-pudding" Model. But this model was not satisfied and discarded.

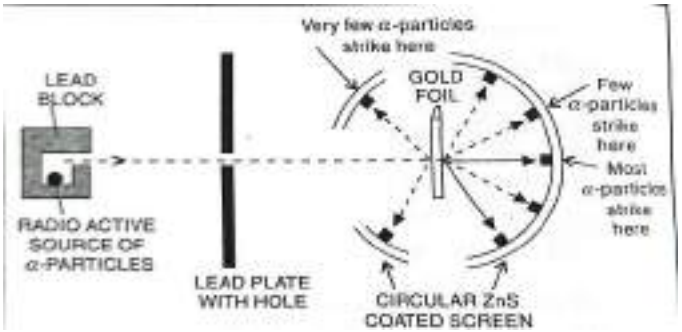
RUTHERFORD’S ATOMIC MODEL:

This model is also known as Rutherford's Gold-foil Experiment/Rutherford’s α-scattering Experiment. ***In 1911,*** **E.** ***Rutherford*** gave the first information about the most accurate model of an atom. Rutherford’s gold foil experiment can be studied in two steps.

1. Experiment
2. Observation & Conclusion 3.

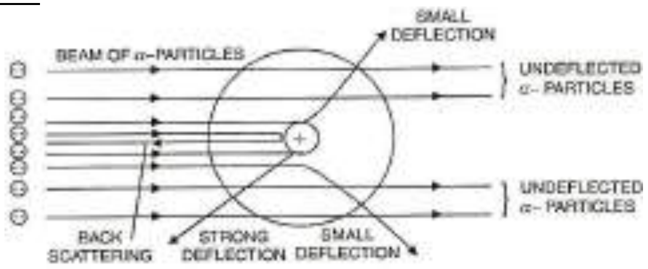
1. EXPERIMENT:

Rutherford bombarded a thin sheet of gold foil (having thickness0.00004 cm) with α-particles (He^{2+}) coming out of radioactive material Radium. A thin lead plate with a hole in it was placed to form a beam in it. A circular screen coated with zinc sulphide in the inner side was placed on the other side of the foil as shown in fig.2.



(Fig.2 Gold foil experiment)

2. OBSERVATION & CONCLUSION:



(Fig.3 scattering of α – particles in Gold foil experiment)

1. Most of the α-particles passed straight through the gold foil without any deviation. This clearly indicates that most of the parts of an atom are empty.
2. Some α-particles were found to be deflected in angle. This indicatesthe presence of a massive positively charged particle at the center. This heavy positively charged body is called nucleus.
3. A very small number of alpha particles were retraced its path (about 1 in 10,000) as shown in fig.3.

POSTULATES OF RUTHER FORD'S ATOMIC MODEL:

Based on the conclusions drawn from the α-scattering experiment, Rutherford proposed an atomic model as follows:

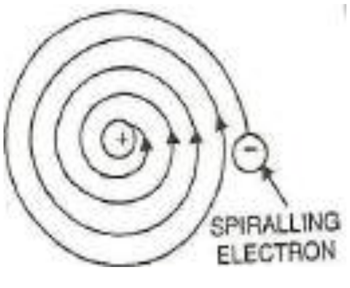
- An atom consists of two parts; they are (i) Nucleus and
  (ii) extra nuclear part.
- Every atom consists of a massive positively charged centre, called nucleus.
- The whole mass of an atom is concentrated at the nucleus.
- Electrons having negative charge revolve around the nucleus in a circular path known as orbit with a high speed.
- The revolving of electron around nucleus is balanced by the electrostatic force of attraction i.e centripetal force (acting inward) and centrifugal force (acting outward). The Rutherford’s atomic model is compared the Solar system.

DRAWBACKS OR FAILURES OF RUTHERFORD'S ATOMIC MODEL:

1. Stability of Atom:

The theory fails to explain the stability of atoms. According to Clark Maxwell when a charged particle revolves around another charged particle, the revolving charged particle emits (loses) energy continuously.

As a result of which the electron should lose its energy continuously and should be attracted towards the nucleus forming a coil like structure and ultimately fall into the nucleus as shown in fig.4. However this never happens. Thus Rutherford’s could not explain the reason.



2. The theory fails to explain about the discontinuity of atomic spectra.

3. It fails to explain the arrangement of electron around (fig.4) nucleus in space.

Atomic Number (Z):

The total number of protons present in an atom is called its atomic number.

Atomic Number = No. of Protons (p)
= No of electron (e-)

Examples:

Element	No of Protons (p)	Atomic No. (Z)
Carbon (C)	6	6
Sodium (Na)	11	11
Iron(Fe)	26	26

Mass Number (A): The total number of protons and neutrons present in one atom of an element is called its mass number.

Mass Number (A) = No. of Protons (p) + No. of Neutrons (n)

Element	No of Protons (p)	No. of Neutrons (n)	Mass No. (A)
Carbon (C)	6	6	12
Oxygen (O)	8	8	16
Sodium (Na)	11	12	23

Isotopes:

Isotopes are the atoms of the same element having the same atomic number but different mass numbers.

Examples: Hydrogen has three isotopes: Protuium ($^{1}_{1}H$), Deutoriium ($^{2}_{1}H$) and Tritiium ($^{3}_{1}H$)
$^{35}_{17}Cl$ and $^{37}_{17}Cl$

Properties of Isotopes:

- These are atoms of the same element.
- These have similar chemical properties.
- These have different physical properties such as m.pt., b.pt, density, viscosity, etc.

Isobars:

Isobars are the atoms of different elements having the same mass number but different atomic numbers.

Examples: $^{40}_{18}Ar$ and $^{40}_{20}Ca$

Properties of Isotopes:

- These are atoms of different elements.
- These have different chemical properties.
- These have different physical properties such as m.pt., b.pt, density, viscosity, etc.

Isotones:

Isotones are the atoms of the elements having the same no of neutron.

Examples: $^{23}_{11}Na$ and $^{24}_{12}Mg$ $^{27}_{13}Al$ and $^{28}_{14}Si$ $^{31}_{15}P$ and $^{32}_{16}S$

Note: The subatomic particles (protons and neutrons) present inside the nucleus are collectively called as nucleons.

BOHR'S ATOMIC MODEL /(BOHR'S POSTULATES):

Neils Bohr, a Danish physicist in 1913 put forward a theory also known as Bohr's postulates. He modified Rutherford's atomic model with the help of Planck's Quantum Theory.

The main postulates of Bohrs atomic model are

1. Every atom consists of a heavy positively charged body at the centre called 'nucleus' andelectrons revolve around the nucleus in a fixed circular paths called 'shells', 'orbits' without radiating energy . So these orbits are also known as 'stationary states'.

2. The stationary states are associated with some fixed amount of energy and hence they are also known as energy level. The energy level are designated as **K, L, M, N, O......etc.** and numbered as n=1, 2, 3, 4, 5.....etc respectively.

3. The energy content of various shells are compared by the formula:

$$E_n = -\frac{2\pi^2 me^4 z^2}{n^2 h^2}$$

where, *m = mass of an electron*

e = charge of electron

z = Atomic number

n = No. of shell or principal quantum number (1,2,3.......)

h = Planck's constant = 6.625 x 10^{-27} erg sec.

$$E_n = -\frac{1312}{n^2} \text{ KJ / Mol,} \qquad \text{(for Hydrogen atom)}$$

4. The energy content increases on moving from lower to higher shells and become zero for the shell which is present at an infinite distance from the nucleus.

5. The energy levels are not equally spaced.

6. Only those circular paths are allowed for the electrons to revolve for which the **angular momentum** value (mvr) is a whole number multiple of $\frac{h}{2\pi}$, that is

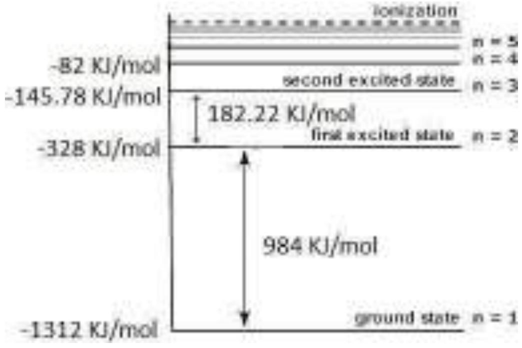
$$mvr = n\frac{h}{2\pi}, \text{ where } n = 1, 2, 3,..........$$

Here, m = mass of the electron,

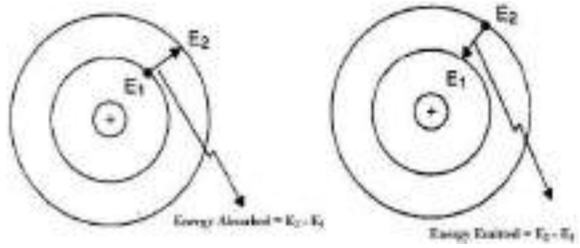
v = tangential velocity of the revolving electron

r = radius of the orbit. (fig.5)

h = Planck's constant,



7. When the electrons in an atom are in their normal energy state (ground state), they keep on revolving in their respective orbits without losing energy.



8. When external source of energy is supplied to an atom, the electrons jump from lower ground states to the higher excited states by absorbing energy. Electrons in the excited states are unstable and jump back to the lower ground states by releasing energy. The electron thus absorbed and emits energy in the form of light is known as quanta which is the origin of spectral lines. (Fig.6)

Failures of Bohr’s Atomic Model:

1. It fails to explain the spectra of multi-electron atoms.
2. According to Bohr’s atomic model, the electrons revolve in a two-dimensional space however electrons revolve in three-dimensional space called orbitals.
3. It fails to explain the relative intensities of spectral lines.
4. It fails to explain the splitting up of spectral lines when exposed to electric field (Stark Effect) and magnetic field (Zeeman Effect).
5. It fails to explain the cause of chemical combinations.
6. One of the major drawback is that, according to Bohr the position and velocity of an electron can be calculated accurately. However according to Heisenberg’s uncertainty principle it is impossible to measure simultaneously the position and momentum.

BOHR-BURY SCHEME:

Bohr-Burry scheme deals with the arrangement of electrons in various shells. Various postulates of the scheme are:

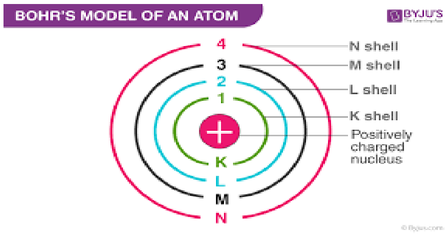
1. A shell can contain a maximum $2n^2$ number of electrons as in fig.7. Where n = number of the shell.

Shell	n	Maximum No. of Electrons = $2n^2$
K	1	$2 X 1^2 = 2$
L	2	$2 x 2^2 = 8$
M	3	$2 x 3^3 = 18$
N	4	$2 X 4^4 = 32$, and so on.

2. The outer most shell (valence shell) of an element cannot hold more than ‘8’ electrons.

3. The penultimate shell (the shell just before the outer most shell) cannot hold more than ‘18’ electrons.

4. A higher orbit may start filling before the lower orbit is completely filled.



(Fig.7)

AUF-BAU PRINCIPLE:

Aufbau principle may be stated as electrons are filled in various orbital in order of their increasing energy.

The sub-shell with lowest energy is filled with electron first and those with higher energies are filled later. The energy content of the various sub-shells can be compared by (n+ l) rule.

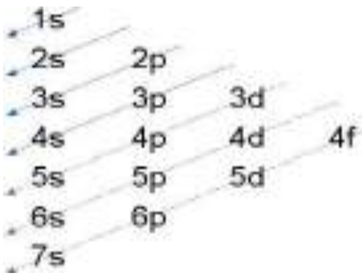
n = principal quantum number

l = azimuthal quantum number

Sub-shell→	**1s**	**2s**	**2p**	**3s**	**3p**	**3d**	**4s**	**4p**	**4d**	**4f**	**5s**	**5p**
(n+l)	1+0 =1	2+0 =2	2+1 =3	3+0 =3	3+1 =4	3+2 =5	4+0 =4	4+1 =5	4+2 =6	4+3 =7	5+0 =5	5+1 =6

Hence the increasing order of energy content of sub-shells is

1s < 2s < 2p < 3s < 3p <4s <3d < 4p

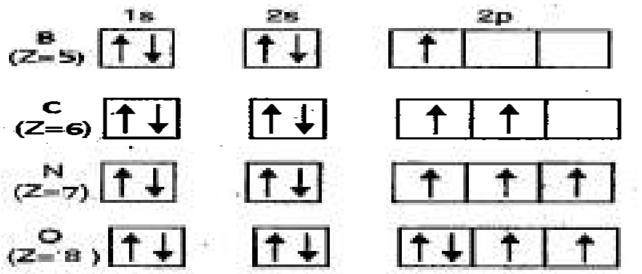


HUND'S RULE:

Hund's rule may be stated as no electrons take place in the p, d, and f-sub shells until each degenerate orbital contains one electron."

Let us consider the electronic configurations of the following elements.

In case of boron the 5th electron is occupied by the $2p_x$ orbital. In carbon the 6th electron will not be paired with the electron of the $2p_x$ orbital, rather it will be occupied by the $2p_y$ orbital. Similarly in case of nitrogen all the $2p$ – electrons will remain unpaired. The rule is also called 'maximum multiplicity rule' because the total spin value of all the electrons of degenerate orbitals of a given sub-shell becomes maximum if they are arranged as per Hund's rule.



NOTE :- Degenerate orbitals are the orbitals having same or nearly same energy content. For example $2p_x$, $2p_y$, $2p_z$ are degenerate orbitals.

ELECTRONIC CONFIGURATIONS:-

Electronic configuration is the arrangement of electrons of an atom in different sub-shells/orbitals in the increasing order of their energy . The electronic configurations of some elements are given below:

Elements	Electronic configurations	Elements	Electronic configurations
${}_{1}\mathrm{H}$	$1s^1$	${}_{16}\mathrm{S}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^4$
${}_{2}\mathrm{He}$	$1s^2$	${}_{17}\mathrm{Cl}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^5$
${}_{3}\mathrm{Li}$	$1s^2\ 2s^1$	${}_{18}\mathrm{Ar}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^6$
${}_{4}\mathrm{Be}$	$1s^2\ 2s^2$	${}_{19}\mathrm{K}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^6 4s^1$
${}_{5}\mathrm{B}$	$1s^2\ 2s^2 2p^1$	${}_{20}\mathrm{Ca}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^6 4s^2$
${}_{6}\mathrm{C}$	$1s^2\ 2s^2 2p^2$	${}_{21}\mathrm{Sc}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^6 4s^2 3d^1$
${}_{7}\mathrm{N}$	$1s^2\ 2s^2 2p^3$	${}_{22}\mathrm{Ti}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^6 4s^2 3d^2$
${}_{8}\mathrm{O}$	$1s^2\ 2s^2 2p^4$	${}_{23}\mathrm{V}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^6 4s^2 3d^3$
${}_{9}\mathrm{F}$	$1s^2\ 2s^2 2p^5$	**${}_{24}\mathrm{Cr}$**	**$1s^2\ 2s^2 2p^6\ 3s^2 3p^6 4s^1 3d^5$**
${}_{10}\mathrm{Ne}$	$1s^2\ 2s^2 2p^6$	${}_{25}\mathrm{Mn}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^6 4s^2 3d^5$
${}_{11}\mathrm{Na}$	$1s^2\ 2s^2 2p^6\ 3s^1$	${}_{26}\mathrm{Fe}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^6 4s^2 3d^6$
${}_{12}\mathrm{Mg}$	$1s^2\ 2s^2 2p^6\ 3s^2$	${}_{27}\mathrm{Co}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^6 4s^2 3d^7$
${}_{13}\mathrm{Al}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^1$	${}_{28}\mathrm{Ni}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^6 4s^2 3d^8$
${}_{14}\mathrm{Si}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^2$	**${}_{29}\mathrm{Cu}$**	**$1s^2\ 2s^2 2p^6\ 3s^2 3p^6 4s^1 3d^{10}$**
${}_{15}\mathrm{P}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^3$	${}_{30}\mathrm{Zn}$	$1s^2\ 2s^2 2p^6\ 3s^2 3p^6 4s^2 3d^{10}$

Exceptional Case:

Some elements like ‘Cr’ and ‘Cu’ show exceptional electronic configurations.

The electronic configurations of ‘Cr’ & ‘Cu’ should be

${}_{24}\mathrm{Cr}$ = [Ar] $4s^2 3d^4$ & ${}_{29}\mathrm{Cu}$ = [Ar] $4s^2 3d^9$ respectively.

But the actual electronic configurations are

${}_{24}\mathrm{Cr}$ = [Ar] $4s^1 3d^5$ & ${}_{29}\mathrm{Cu}$ = [Ar] $4s^1 3d^{10}$

The exceptional electronic configuration is due to the fact that half-filled and ful-filled orbitals are more stable due to the orbital symmetry and exchange energy.

ASSIGNMENT

SECTION -A

SHORT QUESTIONS (2 MARKS)

1. (i) What is the mass and charge of electron?

(ii) What is the mass and charge of proton?

(iii) Calculate the number of electron, proton and neutron of Al^{3+}, Na^{+}, O^{2-}, Mg^{2+}, Ca^{2+}

(iv) In Rutherford's Gold foil experiment alpha particles are nothing but what?

(v) Maximum number of electron present in valence shell and penultimate shell is equal to __and___.

(vi) When electron gain some energy it jumps from______ to ______. When it lose energy it jumps from_____ to _____.

(vii) Define Hund's rule.

(viii) Define Auf-Bau principle.

(ix) Define Isotope and give an example it.

(x) Define Isobar and give an example it.

(xi) Define Isotone and give an example it.

(xii) The electron absorb or emit energy in the form of ______.

SECTION -B

LONG QUESTIONS (4 MARKS)

2(i) What are the drawbacks of Rutherford's Gold foil experiment?

(ii) Explain the Bohr-Bury scheme.

(iii) Define (n+l) rule. 3d orbital has more energy than 4s. Explain.

SECTION – C

LONG QUESTIONS (6 MARKS)

3(i) Write the observation and conclusion of Rutherford's Gold foil experiment.

(ii) Write the postulates of Bohr's atomic model.

__________*_________

CHAPTER-2

(CHEMICAL BONDING)

CHEMICAL BONDING

LEARNING OBJECTIVES

- DEFINITION
- TYPES OF CHEMICAL BOND WITH EXAPMLES

CHEMICAL BOND:

It is defined as the force of attraction which holds together the constituent atoms in a molecule or ion is called chemical bond.

Types of Chemical Bonding:

Depending upon the types of bond formation (transfer or sharing of electrons), chemical bonding may be classified into the following types:

1. Ionic Bonding or Electrovalent bonding
2. Covalent bonding
3. Co-ordinate bonding or Dative Bonding

IONIC OR ELECTROVALENT BONDING:

The chemical bond which is formed by the complete transfer of one or more valance electrons from one atom to another is called ionic or electrovalent bond and the compound formed is called ionic compound or electrovalent compound.

The formation of ionic bond involves formation of a positively charged cation by loss of electron and Formation of a negatively charged anion by gain of electron. Electrostatic force of attraction between the oppositely charged ions helps to combine the ion to form a molecule.

Example: I: Formation of NaCl.

The electronic configurations of Na and Cl are given below:

$_{11}$Na – $1s^2 2s^2 2p^6 3s^1$
$_{17}$C*l* – $1s^2 2s^2 2p^6 3s^2 3p^5$

The electronic configurations indicate the presence of one and seven valence electrons in sodium and chlorine respectively. During the formation of NaCl, the sodium atom donates its valence electron completely to the chlorine atom. Na becomes Na^+ with 8 electrons in valence shell and attains the nearest noble gas configuration of Ne; while C*l* atom becomes Cl^- with '8' electrons in valence shell and attains the nearest noble gas configuration of Ar.

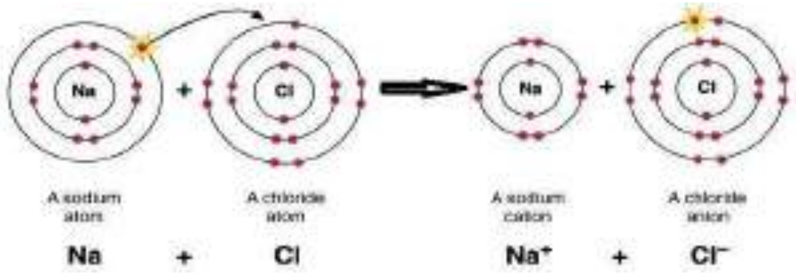
$_{11}Na \; - \; e^- \rightarrow Na^+$

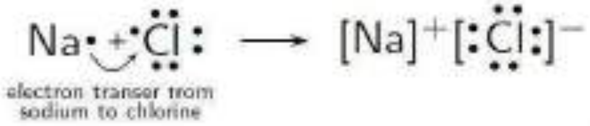
(2, 8, 1) (2, 8)

$_{17}Cl \; + \; e^- \rightarrow Cl^-$

(2, 8, 7) (2, 8, 8)

(Electrostatic Force of Attraction) $\rightarrow$ $Na^+ \, Cl^-$ or NaCl





Now the electrostatic force of attraction between the oppositely charged ions Na^+ and Cl^- results in the formation of **NaCl**.

Example :II : Formation of $MgCl_2$

The electronic configurations of Mg and Cl are given below:

$_{12}Mg$ - $1s^2 2s^2 2p^6 3s^2$

$_{17}Cl$ - $1s^2 2s^2 2p^6 3s^2 3p^5$

An atom of magnesium and chlorine contain 2 and 7 numbers of valence electrons respectively. During the formation of $MgCl_2$, one of the valence electrons of Mg is completely transferred to one atom of chlorine while the second valence electron is transferred to another atom. Now Mg becomes **Mg^{2+}** with **8** valence electrons, with electronic configuration similar to that of Neon. On the other hand each chlorine atom after accepting an electron becomes Cl^- ion with 8 valence electrons, acquiring electronic configuration similar to that of Argon.

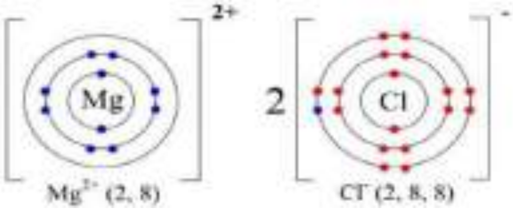
$_{12}Mg \; - 2e^- \rightarrow Mg^{2+}$

(2, 8, 2) (2, 8)

$2\,Cl \; + \; 2e^- \rightarrow 2\,Cl^-$

(2, 8, 7) (2, 8, 8)

(Electrostatic Force of Attraction) $\rightarrow$ $Mg^{2+} \; 2Cl^-$ or $MgCl_2$



The electrostatic force of attraction between calcium and chloride ions results in the formation of **$MgCl_2$.**

Other examples of ionic compounds: KCl, KBr, KI, CaF_2, $CaBr_2$, CaI_2, Na_2O, K_2O, CaO, MgO, Na_2S, K_2S etc.

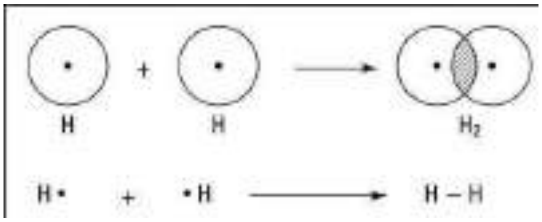
Characteristics of ionic compounds***:*** Ionic compounds possess the following characteristic.

- These exist in solid states.
- These are hard and rigid.
- These are crystalline in nature.
- These have high melting and boiling point.
- These have high densities.
- These are soluble in polar solvents like water, but are insoluble in non-polar solvents like CCl_4,ethers, benzene, toluene [$C_6H_5 - CH_3$], petrol, diesel, kerosene, etc.
- These are bad conductor of electricity in solid states, but are good conductor in molten, fusedor solution state.
- These do not show isomerism.
- These are polar in nature.

COVALENT BOND:

The chemical bond formed by the mutual (equal) sharing of valence electrons between two atoms is called covalent bond and the compound formed is called covalent compound.

A covalent bond may be formed between the atoms similar or dissimilar elements. When two, four and six electrons are shared between two atoms, then a single, double and a triple bond are formed respectively.



Example : **I** ***:*** Formation of H_2 molecule.

The electronic configuration of ‘H’ is

$_1H - 1s^1$

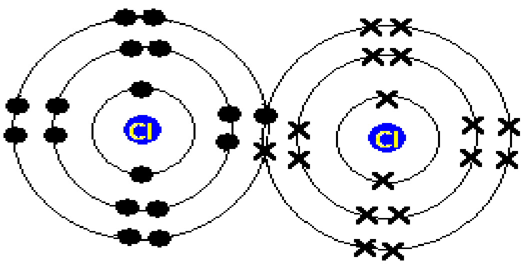
The electronic configuration indicates the presence of 1 valence electron in ‘H’ and requires one more electron to become duplet by acquiring the configuration of He. Thus, each hydrogen atom shares its electron with each other to form a covalent bond.

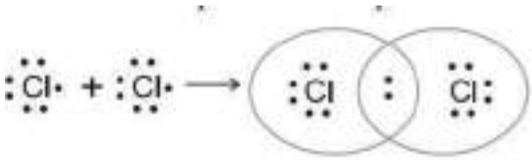
Example : 2 : Formation of Cl_2 molecule.

The electronic configuration of 'Cl' is

$_{17}Cl - 1s^2 2s^2 2p^6 3s^2 3p^5$

The electronic configuration indicates the presence of 7 valence electrons in ‘Cl’ and requires one more electron to become octet. Thus, each chlorine atom shares one of its valence electron with each other to form a covalent bond.



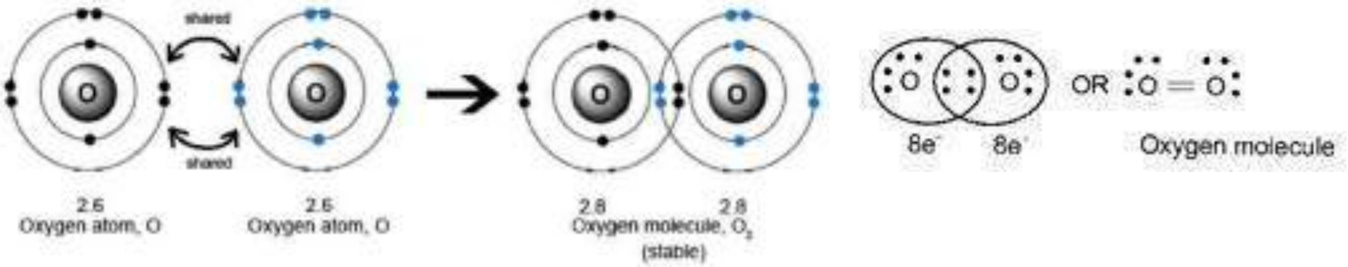


Example :3 **:** Formation of O_2 molecule.

The electronic configuration of 'O' is

$_{8}O$- $1s^2 2s^2 2p^4$

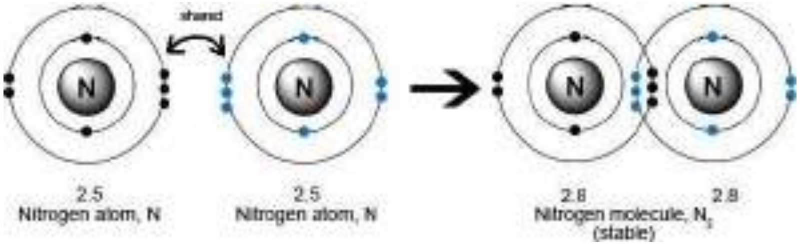
The electronic configuration indicates the presence of 6 valence electrons in ‘O’ and requires two more electrons to become octet. Thus, each oxygen atom shares two of its valence electrons with each other to form a double covalent bond.

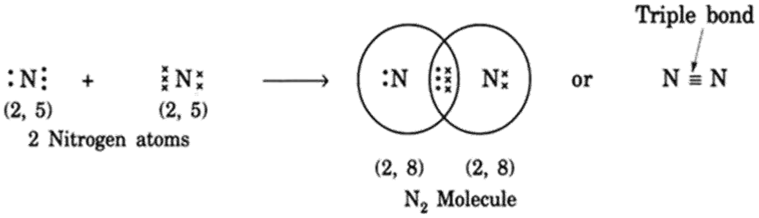


Example : **4 :** Formation of N2 molecule.The electronic configuration of 'N' is

$_{7}N$- $1s^2 2s^2 2p^3$

The electronic configuration indicates the presence of 5 valence electrons in ‘N’ and requires three more electrons to become octet. Thus, each nitrogen atom shares three of its valence electrons with each other to form a triple covalent bond.





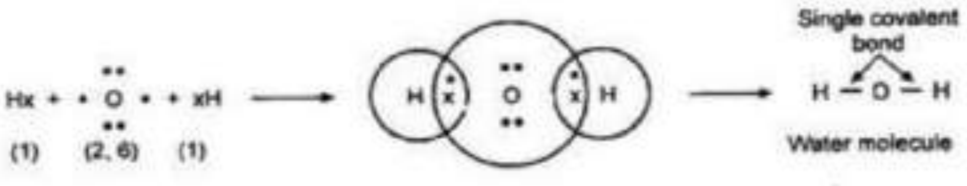
Example : 5 Formation of water molecule(H_2O)

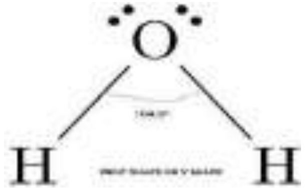
The electronic configuration of 'O' and ‘H’ are:

$_{8}O$- $1s^2 2s^2 2p^4$

$_{1}H$ - $1s^1$

The electronic configuration indicate the presence of 6 and 1 valence electrons in ‘O’ and ‘H’ respectively. The central ‘O’ atom requires two more electrons to become octet while each hydrogen atom needs one electron to become duplet. Thus, each hydrogen atom shares its valence electron with the valence electrons of oxygen to form covalent bonds. The shape of water molecule is bent shape or ‘V’ shape with bond angle 104.5^0.



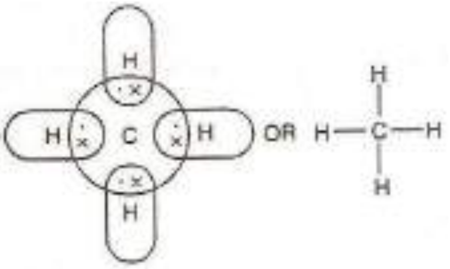


Example: 6: Formation of methane (CH_4) molecule.

The electronic configuration of 'C' and 'H' are

$_6$C - $1s^2 2s^2 2p^2$

$_1$H - $1s^1$



The electronic configurations indicate the presence of 4 & 1 valence electrons in ‘C’ & ‘H’ respectively. Thus, the central carbon atom requires four more electrons to become octet and ‘H’ requires 1 more electron to become duplet.

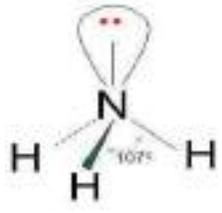
Thus, each hydrogen atom shares its electron with one valence electron of carbon to form four single covalent bonds. By sharing an electron each hydrogen atom becomes duplet while carbon becomes octet. The shape of CH_4 is tetrahedral with a bond angle of $109^0 28'$.

Example: 7: Formation of ammonia (NH_3) molecule.

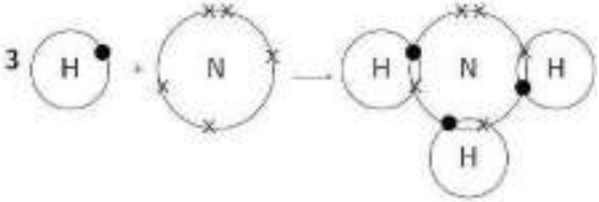
The electronic configuration of 'C' and 'H' are

$_7$N - $1s^2 2s^2 2p^3$

$_1$H - $1s^1$



Pyramidal



The electronic configurations indicate the presence of 5 & 1 valence electrons in ‘N’ & ‘H’ respectively. Thus, the central nitrogen atom requires three more electrons to become octet and each ‘H’ atom requires 1 more electron to become duplet.

Thus, each hydrogen atom shares its valence electron with one valence electron of nitrogen to form three single covalent bonds. The shape of NH_3 is pyramidal with a bond angle of 107^0.

Other examples of covalent compounds: , F_2, Br_2, I_2, BF_3, $AlCl_3$, HCl, HF, SiO_2, etc.

Characteristics of Covalent Compounds:

- These may exist in all the three states of matter i.e. solid, liquid and gaseous states.
- These are generally soft.
- These are non-crystalline in nature.
- These are generally insoluble in polar solvent like water, but are soluble in non-polar solvents such as CCl_4, ethers, benzene, toluene, petrol, diesel, kerosene, etc.
- These have low melting and boiling points.
- These have low densities.
- These are bad conductor of heat and electricity.
- These compounds may show isomerism.
- These are generally non-polar.

CO-ORDINATE BOND:

The chemical bond formed by the partial sharing of a lone pair ofelectrons between two atoms is called a co-ordinate or dative bond.

Conditions for the formation of co-ordinate or Dative bond:

- One of the participating atoms should have at least one lone or unshared pair of electrons.
- The other atom should be in short of a pair of electrons than the nearest inert gas.

A co-ordinate bond is represented by an arrow (→) sign, the head of which points towards the acceptor atom, while the tail points towards the donor atom.

Since co-ordinate bond has some polar character, it is also known as dative or semi-polar bond or co-ionic bond.

EXAMPLE 1:

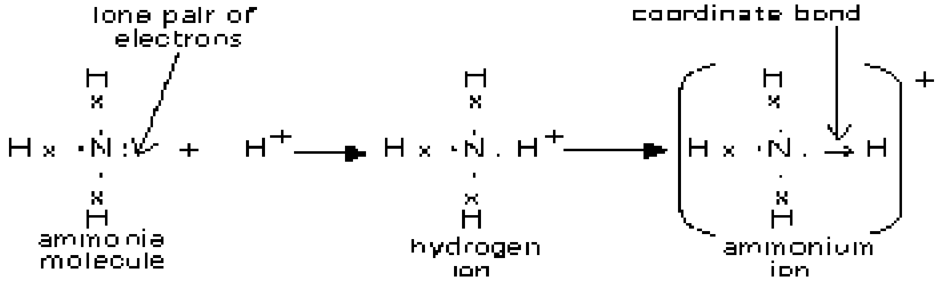
Formation of ammonium ion (NH_4^+).

Ammonium ion (NH_4) is formed by the combination of NH_3 and H^+ ion.

$NH_3 + H^+ \rightarrow NH_4^+$

Ammonia (NH_3) contains a lone pair of electrons over 'N' while 'H^+' ion contains no electron and requires two electrons to become stable.

Thus the lone pair of electrons of nitrogen in NH_3 is partially shared with H^+ ion and a Co-ordinate bond is formed.

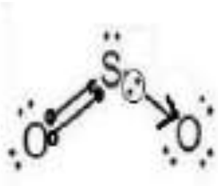


Example: 2: Formation of Sulphur dioxide (SO_2) molecule.
The electronic configurations of 'S' and 'O' are

$_{16}S$ - $1s^2 2s^2 2p^6 3s^2 3p^4$

$_8O$- $1s^2 2s^2 2p^4$

The electronic configurations indicate the presence of 6 valence electrons in each of 'S' & 'O'. the central sulphur atom forms a double covalent by sharing two of its valence electrons with two valence electrons of one of the oxygen atoms. One of the lone pair of electrons of the central sulphur atom is partially shared with another oxygen atom to form a co-ordinate bond.



Characteristics of Coordinate Compound:

- These bonds are rigid and directional.
- Coordinate bonds do not ionise in a state of fusion of solution.
- They are usually insoluble in water but dissolve in non-polar solvents.
- Their melting and boiling points are higher than those of covalent compounds and lowerthan those of ionic compounds.
- These are semi-polar, that is more polar than covalent compounds and less polar than ioniccompounds.
- They show isomerism.

ASSIGNMENT

SECTION -A

SHORT QUESTIONS (2 MARKS)

1. (i) Define chemical bonding?

(ii) Define valency? What is the valency of Fe in FeO?

(iii) Give two examples of monoatomic divalent radical.

(iv) Give two examples of monoatomic tetravalent radical.

(v) Define ionic bonding. Give an example of it

(vi) Define covalent bonding. Give an example of it

(vii) Define co-ordinate bonding. Give an example of it

(viii) Give two examples having a co-ordinate bond and covalent bond in it.

(ix) An atom having number of proton equal to 13, what will be the number of neutron in it?

(x) How many valence electrons are there in an atom having number of electron equal to 19 and 26 .

SECTION -B

LONG QUESTIONS (4 MARKS/6 MARKS)

2 (i) Define ionic bonding. Explain the formation of NaCl and $CaCl_2$

(ii) Define Co-valent bonding. Explain the formation of CH_4 and H_2O.

(iii) Define Co-Ordinate bonding. Explain the formation of SO_2, H_3O^+, NH_4^+

__________*_________

CHAPTER- 3

(ACID-BASE THEORY)

ACID-BASE THEORY

LEARNING OBJECTIVES

- ARRHENIUS THEORY OF ACID AND BASE
- BRONSTED –LOWRY THEORY OF ACID AND BASE
- LEWIS THEORY OF ACID AND BASE
- NEUTRALISATION OF ACID AND BASE
- DEFINITION OF SALT AND TYPES

There are many practical methods to identify the acidic and basic character of substances which include use of p^H meter, indicator, litmus paper, etc. However various theory were put forward by different workers most widely accepted theories are:

1. Arrhenius Theory
2. Bronsted- Lowery Theory
3. Lewis Theory

1. ARRHENIUS THEORY:

(i) According to Arrhenius theory, Acids are the substances which produce **H^+ ions** in aqueous solution while bases are the substances which produce **OH^- ions** in aqueous solution.

Example of Acid: $HCl \xrightarrow{water} H^+(aq) + Cl^-(aq)$

Other examples of acids are: HNO_3, H_2SO_4, CH_3COOH, etc.

Examples of Bases: $NaOH_{(s)} \xrightarrow{water} Na^+(aq) + OH^-(aq)$

Other examples of bases are: LiOH, KOH, NH_4OH, $Ca(OH)_2$, $Al(OH)_3$, etc

(ii)According to Arrhenius theory acidic character of a substance is due to the formation of H^+ ion and basic character are due to presence of OH^- .

(iii) Strong acids and bases are the ones which dissociate completely in a solution. For, example HCl is a strong acid and dissociates completely into H+ and Cl- ions in the solution. Similarly, NaOH dissociates completely in solution. Similarly weak acids and bases are the ones which do not dissociate completely in a solution. For, example CH3COOH is a weak acid and NH4OH is a weak base. They do not dissociate completely in solution.

(iv) According to Arrhenius theory an acid reacts with a base to form salt and water and the reaction is known as neutralization reaction.

$$\underset{(\text{ Acid })}{HCl} + \underset{(\text{Base})}{NaOH} \rightarrow \underset{(\text{Salt})}{NaCl} + \underset{(\text{Water})}{H_2O}$$

Neutralization reaction may be represented as:

$$\underset{(\textit{from acid})}{H^+(aq)} + \underset{(\textit{from base})}{OH^-(aq)} \rightarrow H_2O_{(l)}$$

Limitations Arrhenius Theory:

i. According to Arrhenius theory, Acids are the substances which produce **H^+ ions** in aqueous solution . However H^+ ion does not exist independently in aqueous solution. It combines with H_2O to give hydronium ion (H_3O^+).

$$H^+ + H_2O \rightarrow H_3O^+$$

ii. The theory fails to explain the acidic and basic nature of the substances in solvents other than water.

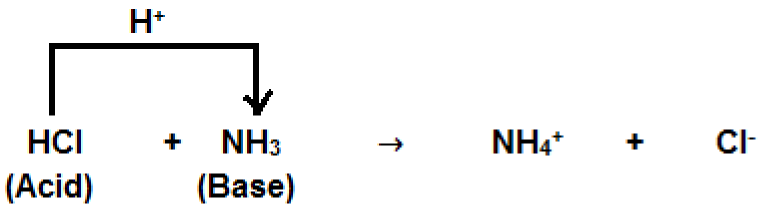
iii. The theory fails to explain the acidic nature of the substances like SO_2, CO_2, SiO_2, P_2O_5, BF_3, $AlCl_3$, etc. which cannot provide H^+ ions.

iv. It could not explain the neutralization reaction in absence of water other than solvent like water.

2. BRONSTED –LOWRY THEORY OF ACID AND BASE:

(i) According to Bronsted-Lowry theory Acids are the substances which give a proton ***(H^+ ion)*** to any other substance, while bases are the substances which accept a proton ***(H^+ ion)*** from any other substances .

In other words, acids are proton donors whereas bases are proton acceptors. For example:



Since, HCl has donated a proton (to NH_3), it acts as an acid**.** On the other hand, NH_3 has accepted a proton from HCl and thus it acts as a base.

Other examples of Acids are:

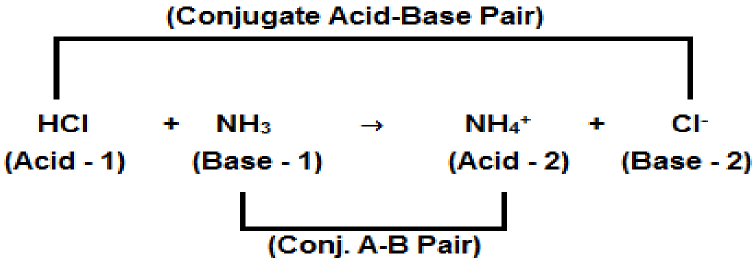
- All Arrhenius acids are Bronsted-Lowry acids; (HCl, HNO_3, H_2SO_4, H_3PO_4, CH_3COOH, H_2CO_3 etc.), however the reverse is not true.
- Ions having capacity to donate H^+ ion: (HS^-, HCO_3^-, H_4PO^{2-}, H_4S^- etc.)

Examples of Bases:

- Neutural molecules such as: H_2O, NH_3, RNH_2, PH_3, AsH_3, etc.
- Ions having capacity to accept *H^+ ion,* like OH^-, CN^-, HCO_3^-, etc.

(ii)According to this theory an acid reacts with a base to form another pair of acid and base.

i. For example**:**



The pair of acid and base which differ by a proton (H^+ion) is called a conjugate acid-base pair.

Acid $- H^+$ →Conjugate base

Base + H^+ → Conjugate acid

ii. The substances such as H_2O, HS^-, $HCO^{3\,-}$, $HPO_4^{2\,-}$, HSO_4^-, etc. which act as both acid (proton donor) as well as base (proton acceptor) are called amphoteric substances.

iii. Stronger is an acid weaker is its conjugate base and vice versa.

$$\underset{\text{[Strong acid]}}{HCl} + H_2O \rightarrow H_3O^+ + \underset{\text{[Weak base]}}{Cl^-}$$

Limitations of the theory:

i. It fails to explain the acidic nature of the acidic oxide such as SiO_2, CO_2, SO_2, BF_3, etc. whichcannot donate H^+ ion.

ii. It fails to explain the basic nature of the basic oxide such as Na_2O, K_2O, CaO etc. which cannot accept H^+ ion.

iii. It fails to explain the reaction between some acids and bases which do not give another pair of acid and base. Example: $HCl + NaOH \rightarrow NaCl + H_2O$.

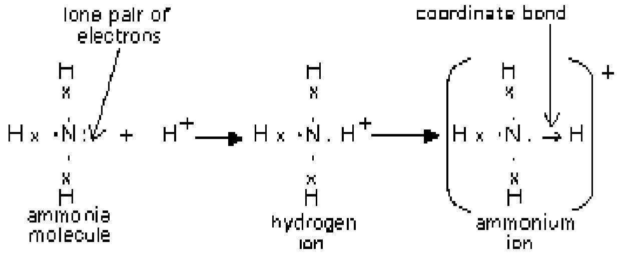
Note: Some conjugate acid-base pairs are given below:

Acid	Conjugate- Base	Base	Conjugate- Acid
HCl	Cl^-	Br^-	HBr
H_2SO_4	HSO_4^-	CN^-	HCN
NH_4^+	NH_3	O^{2-}	OH^-
H_2O	OH^-	NH_3	NH_4^+
		H_2O	H_3O^+

3. LEWIS THEORY: -

(i) According to Lewis theory -Acids are the substances which can accept a pair of electrons from any other substance, while bases are the substances which can donatea pair of electrons to any other substance. In other words, acids are electron acceptors while bases are electron donors.

(ii) According to this theory, an acid reacts with a base to form a co-ordinate or dative bond. For example, the reaction between NH_3 (Lewis base) and H^+ (Lewis acid) results in the formation of a dative bond.



Examples of acids: -

i. Cations like: CH_3^+, H^+, etc.

ii. Neutral molecules containing electron deficient atoms are Lewis acids. For example:BF_3, $AlCl_3$, $FeCl_3$, $ZnCl_2$, etc.

iii. Neutral molecules containing vacant d-orbitals in the central atom for the accommodation ofincoming electrons act as Lewis acids. For example: SiF_4, $SiCl_4$, etc.

iv. The molecules having multiple bonding (= or ≡) between the atoms of different elements are acidic in nature. For example: CO_2 (O = C = O), SO_2, etc.

Examples of Bases: -

i. All anions are Lewis bases: F^-, Cl^-, CO_3^{2-}, etc.

ii. Neutral molecules containing, at least one lone pair of electrons are Lewis bases: Examples: NH_3, PH_3, H_2O, etc.

Limitations:

i. According to this theory, the reaction between an acid and base results in the formation of a dative bond. Formation of a coordinate bond is a slow process. While the reactions between the acids and the bases are instantaneous or fast.

ii. The theory fails to explain the relative strengths of different acids and bases.

iii. It fails to explain reaction between some acids and bases where no coordinate bond is formed.

iv. It fails to explain the acidic nature of well-known acids like HCl, HNO_3, H_2SO_4, etc. which cannot accept electrons.

v. It fails to explain the basic nature of well-known bases like NaOH, KOH, etc. which cannot donate electrons.

vi. It fails to explain acid-catalyzed reactions, where H^+ ion plays important role.

Neutralization of Acids and Bases:

When an aqueous solution of acid is added to aqueous solution of base, a chemical reaction occurs resulting in the formation of a salt and water. This process is called acid-base neutralization reaction.

1. **Neutralization between a Strong Acid and a Strong Base:** A strong acid reacts with a strong base to form a simple or normal salt. Its aqueous solution has a p^H of about 7 and is neutral.

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

2. **Neutralization between a Strong Acid and a weak Base:** A strong acid reacts with a weak baseto form a acidic salt. Its aqueous solution has a $p^H < 7$ and the solution is acidic.

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

3. **Neutralization between a Weak Acid and a Strong Base:** A weak acid reacts with a strong baseto form a basic salt. Its aqueous solution has a $p^H > 7$ and is alkaline.

   Example: $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

4. **Neutralization between a Weak Acid and a Weak Base:** A weak acid reacts with a weak base to form a neutral salt. Its aqueous solution has a $p^H > 7$ and is alkaline.

   Example: $CH_3COOH + NH_4OH \rightarrow CH_3COONH_4 + H_2O$

SALTS

Salt is defined as the crystalline compound which is formed by the complete neutralization of aqueous solution of strong acid and aqueous solution of strong base.

Or

Salts are ionic compounds which produce cation other than H^+ and anion other than OH^- inaqueous solution.

Or

Salts are the compounds formed by the neutralization reaction between acids and bases.

TYPE OF SALTS:

Salts may be classified into the following types:

1. **Normal salts:** These salts are obtained by the reaction between strong acids and strong bases.

Example:	**Acids**	**Normal salts**
	HCl	$NaCl$, KCl, $CaCl_2$, $MgCl_2$, etc
	HNO_3	$NaNO_3$, KNO_3, $Ca(NO_3)_2$, $Mg(NO_3)_2$,
	etc.H_2SO_4	Na_2SO_4, K_2SO_4, $CaSO_4$, $MgSO_4$, etc.
	H_3PO_4	Na_3PO_4, K_3PO_4, $Ca_3(PO_4)_2$, $Mg_3(PO_4)_2$, etc.

2. **Acidic salts:** The salt obtained by the partial replacement of replaceable hydrogen atoms or polybasic acid is called an acidic salt.

Example:

Acids	**Acidic salts**
H_2SO_4	$NaHSO_4$, $KHSO_4$ etc.
H_3PO_4	NaH_2PO_4, KH_2PO_4, Na_2HPO_4, K_2HPO_4

3. **Basic salt:** These are the salts obtained by the incomplete neutralization of poly acidic bases. Such salts contain one or more ‗OH‘ groups.

   Example: Ca (OH)Cl, Mg(OH)Cl, Zn(OH)Cl, $Al(OH)_2Cl$ etc.

4. Double salts: These are the molecular addition compounds obtained from two simple salts, the ions of which retain their identity in aqueous solution. These are stable in solid state only.

Examples: Potash alam($K_2SO_4.Al_2(SO_4)_3.24H_2O$), Mohr‘s Salt [$FeSO_4.(NH_4)_2SO_4.6H_2O$], carnalite ($KCl. MgCl_2.6H_2O$), etc.

5. Complex Salts: These are the salts obtained by the combination of simple salts.These are stable in solid state as well as solution state.

Examples: $K_4[Fe(CN)_6]$, $[Cu(NH_3)_4]SO_4$, etc.

6. **Mixed Salts:** These are the salts which give either more than one cation or more than one anion in aqueous solution.
   Examples: Bleaching powder $CaOCl_2$; Sodium potassium sulphate $NaKSO_4$, etc.

   $CaOCl_2 \text{------------}\rightarrow Ca^{2+} + OCl^{-} + Cl^{-}$ $NaKSO_4 \text{-------}\rightarrow Na^{+} + K^{+} + SO_4^{2-}$

ASSIGNMENT

SECTION -A

1.SHORT QUESTIONS (2 MARKS)

(i)What do you mean by acid & base according to Arrhenius theory?

(ii) Define strong acid and give an example of it?

(iii) Define weak base and give an example of it?

(iv) Define acid and base according to Bronsted-Lowry concept.

(v) Write down the conjugate acids and conjugate bases of H_2O & NH_3.

(vi)What do you mean by conjugate acid-base pair? Explain with a suitable example.

(vii)What is neutralization reaction? Give an example of it.

(viii)Define salt. How does an acidic salt form?

(ix)Define salt. How does a basic salt form?

(x)What is double salt? Give an example.

(xi)What is mixed salt? Give an example.

(xii)Define acid and base according to Lewis concept.

SECTION -B

2. LONG QUESTIONS (6 MARKS)

(i)Define and explain Arrhenius theory of acids and bases.

(ii)Define and explain Bronsted-Lowery theory of acids and bases.

(iii)Define and explain Lewis theory of acids and bases.

(iv)Explain the limitations of Arrhenius theory.

(v)Explain the limitations of Bronsted-Lowery theory.

(vi)Explain the limitations of Lewis theory.

(vii)Justify that all Arrhenius acids are Bronsted-Lowery acids, but all Arrhenius bases are not Bronsted –Lowery bases.

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