ENGINEERING CHEMISTRY

Unit 1.   PHYSICAL CHEMISTRY

1.0   Symbols:

A symbol is defined as an abbreviation / shorthand representation for the full name of an element.

Ex :   Calcium $\rightarrow$ Ca

Hydrogen $\rightarrow$ H

Aluminum $\rightarrow$ Al

Valency :

It is defined as the number of electrons lost, gained or shared with one atom of an element in order to acquire the stable configuration of nearest inert gas element.

Ex : Valency of 'Na' = 1

Valency of 'Ca' = 2.

Variable Valency :

Elements having more than one valency is known as variable valency.

Elements like iron, copper, tin etc posess variable valency.

The valencies of Iron are 2 and 3.

Copper are l and 2.

Tin are 2 and 4.

Radical or Ion:

It is defined as an atom or group of atoms having (+)ve or (-) ve charge.

Classification of Radicals :

On the basis of charge radicals ore of two types –

a) Electro – positive / Basic radicals / Cations :

Radicals which are positive in character .

Ex – K^+, Ag^+, Pb^{2+}

b) Electro- negative radical / Acid radical/ Anion

Radicals which are negative in character.

Ex – Cl^- , SO_4^{2-}, PO_4^{3-}

Radicals can be classified as follows :

a) Simple radicals

b) Compound radicals

a) Simple radicals : It is a radical which contains one or more atoms of the same element.

Ex – Ag^+, S^{2-}, O^{2-}, O_2^{2-}

b) Compound radicals – It contains two or more atoms of different elements forming a single unit.

Ex – SO_4^{2-}, NO_3^-, PO_4^{3-}

On the basis of magnitude of the charge radicals are of five types :

i)   Monovalent Radical : Radical carrying a single (+ve) / (-ve)charge is called a monovalent radical.

Ex- Na^+, Cl^-, F^-, K^+

ii)   Divalent Radical – Radicals having two units of charge are called divalent radical.   Ex – Cu^{2+}, Mg^{2+}, Fe^{2+}, SO_4^{2-}.

iii) Trivalent Radical : Radicals having three units of (+)v e/ (-) ve charge.

Ex- Al^{3+}, Fe^{3+}, PO_4^{3-}, N^{3-}

iv)   Tetravalent Radical : Radicals having four units of (+ve) / (-ve) charge.

Ex – Sn^{4+}, Pb^{4+}, C^{4-} Ferrocyanide $[Fe (CN)_6]^{4-}$.

v)   Pentavalent Radical : Radicals having five units of (+ve) / (-ve) Charge.

Ex. As^{5+}, Sb^{5+}.

Basic and Acid Radicals :

Electro (+ve) radicals/cations are called a basic radical and an electro(-)ve radicals/anions are called acid radicals.
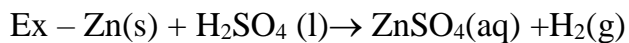
Formula :

The molecule represented with the help of symbols and subscript numbers is known as formula.

Ex – The formula of hydrogen molecule – H_2

,,   hydrogen Chloride – HCl, (Hydrochloric acid)

Chemical Equation :

It is defined as the short hand representation of a chemical reaction with the help of symbols and formulae.

Ex – $Zn(s) + H_2SO_4$ (l)$\rightarrow ZnSO_4(aq) + H_2(g)$

Skeleton Equation :

The equation in which the reactants and the products are involved in a reaction is represented with the help of symbols and formulae is known as skeleton equation.

Ex- $Mg(s) + HCl$ (l) $\rightarrow MgCl_2$ (aq) $+ H_2$ (g)
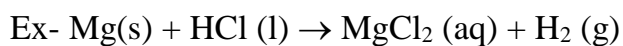
Balanced Equation :

The equation in which the atoms of each kind are equal on both the sides is known as balanced equation.

$Mg(s) + 2HCl$ (l) $\rightarrow MgCl_2(aq) + 2H(g)$.
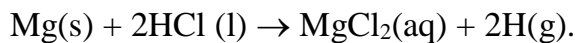
1.1   ATOMIC STRUCTURE :

DISCOVERY OF ELECTRONS:

It is discovered by Thomson' Cathode ray experiment. Charge and mass of electron are 1.602×10^{-19} coulomb and 9.11×10^{-31} kg.

Charge on one mole of electrons $= 96500$ coulombs $= 1F$

DISCOVERY OF PROTONS:

It is discovered by E.Goldstein. The charge on proton is $+1.620 \times 10^{-19}$ coulombs and mass of proton is 1.672×10^{-27}k.g Charge on 1 mole of protons $= 96500$ coulombs.

DISCOVERY OF NEUTRON : (Chadwick)

Chadwick bombarded some lighter elements like Be, B with fast moving α – particles He detected some neutral particles having mass slightly more than the mass of proton. Due to the electrical neutrality of the particles these are called neutrons. This is another fundamental particle present in every atom of the element except proteum. i.e $_1H^1$

mass of neutron $= 1.675 \times 10^{-24}$ gm

$\qquad\qquad = 1.675 \times 10^{-27}$ kg

Atomic number (Z)

It is defined as the no. of unit (+) ve charges or the protons present in the nucleus of an atom.

Atomic number (z)   = Number of protons

$\qquad\qquad\qquad\quad$ = Number of electrons

Ex. : Atomic number of calcium means an atom of calcium consists of 20 protons, 20 neutrons and 20 electrons.

Mass Number (A) :

Mass number is the number of particle that present in the neucleus. i.e mass number (A) = no. of protons + no. of neutrons i.e A = Z + n

e.g – Mass number of 'Na' is 23 and 'Ca' is 40.

Atomic Structure :

Atomic structure means the arrangement of sub-atomic particles (proton, electron, neutron) in an atom. These arrangement are explained through different models.

a) **Thomson's Model**:

According to this model an atom consists of a (+) vely charged sphere within which the (-) vely charged electrons are distributed in arbitrary position in order to neutralize the charge.

Limitation of Thomson's Model :

It fails to explain the experimental facts observed latter. (e.g – Spectral analysis)

Rutherford's Atomic Model:

The atom consists of two sub atomic particles :

i)   Negatively charged particles (electrons)

ii)   Positively charged particles (protons)

Rutherford's Experiment :

Rutherford bombarded α – particles emitted from a radioactive substance on a very thin gold foil. The scattered rays are observed through semicircular screen coated with ZnS salt.
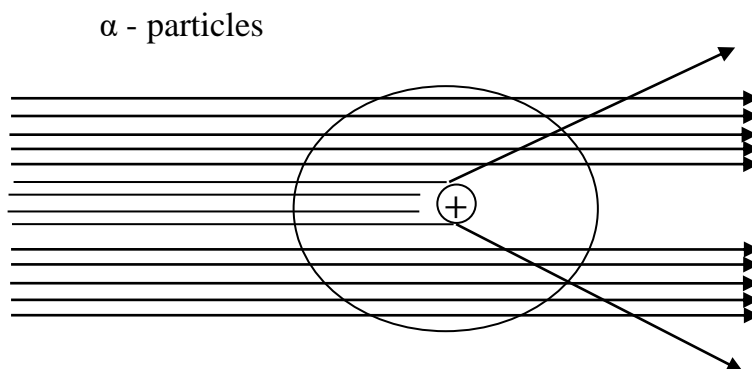
Observation:

i)   A large fraction of α – particles went undeflected i.e they passed straight through the thin foil.

ii)   Few α – particles deflected with certain angles.

iii)   A very few (0.1%) particles even retrace their own path that means reflected through almost 180^0.

Inference:

i)   Since most of the α – particles passed straight through the gold foil without any deflection, a large portion of an atom must be empty.

ii)   Since few α – particles deflected from their normal straight path, there must be some heavy and positively charged body called as nucleus.

Conclusion :

i)   An atom consists of two parts.

a)   Nuclear part b) Extra nuclear part

ii)   Nucleus is a very small but heavy (+)vely charged body, where whole mass of atom is concentrated. This is the central part of the atom.

iii) The extra nuclear part contains the revolving (-)vely charged electron. Those electrons are compared with the planets in the solar system. Thus this atomic model is called planetary model. Electrons are called planetary electrons.

Drawbacks of Rutherford's Atomic model:

i)   If fails to explain the stability of the atom. Because as per Maxwell electron on subsequent revolution may lose energy and may fall in to the nucleus.

ii)   If fails to explain the hydrogen spectra.

iii) The model is silent about the definite energy and velocity posses by the revolving electron.

Bohr's atomic model:

To overcome the limitations of Rutherford Model and to explain the spectral lines, Bohr postulated another model basing on plank's quantum theory.

Postulates of Bohr's atomic Model:

i)   Electrons revolve around the nucleus, the massive central part in certain definite permitted circular path known as stationary states or orbits or shells.

ii)   These shells posses fixed value of energy hence these are called energy levels or stationary states.
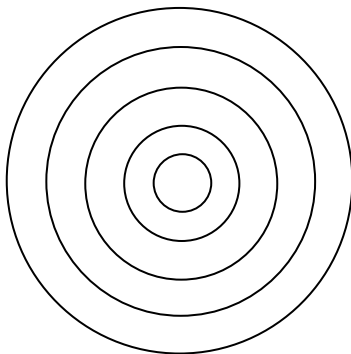
iii)   The energy of an electron in a particular orbit is given by –

$$En = -\frac{313.6}{n^2} KCal/mol \text{ for hydrogen atom}$$

Where the 'n' values are 1, 2, 3…. indicating K, L, M, N …. Shells. Thus the energy in the first shell i.e K-shell = - 313.6 kCal / mole

Energy of the 2nd shell (L-shell) = $\frac{-313.6}{2^2}$ = -78.4 kcal/ mole

Different energy levels aren't equally spaced. It goes on decreasing with the increase in the value of 'n'

iv)   Angular momentum (mvr) of an electron in an orbit is equal to whole number multiple of h/2 .

$$mvr = n\frac{h}{2\pi}$$

Where n = 1, 2, 3, 4…………,

h is plank's constant

   m= mass of the electron

v= Velocity of the electron

r = radius of the orbits.

v)   When electron jumps from higher orbit to lower orbit the energy is released. But when electron jumps from lower orbit to higher orbit energy is absorbed.

The energy absorbed and emitted = $h\nu$ = E $_{higher}$ - E $_{lower}$

Where v is the frequency of energy emitted / absorbed

Thus emission or absorption takes place only when electron jumps from one orbit to another.

vi)　　When the electrons in an atom are in their lowest energy state they keep on revolving in their respective orbits without losing energy. This state of atom is called normal or ground state .After gaining energy electron jumps from lower state to higher state. This state of atom is called excited state which is the unstable state.

Drawbacks of Bohr's atomic Model:

i)　　It fails to explain the line spectra of multi electron system like Be

ii)　　It fails to explain the 3- dimensional model of atom.

iii)　　According to Bohr's model the position and velocity of fast moving small electron can be calculated. But this is against Heisenberg's uncertainty principle.

iv)　　It fails to explain the cause of chemical combination and the shapes of the molecules arising out of it.

v)　　If doesn't explain the splitting of spectral lines on the application of magnetic field.

## BOHR – BURY MODEL:

Postulates are –

1. The maximum number of electrons present in an obit is equal to $2n^2$ where 'n' is the no. of the orbit as follows –

Number of shell (n)		Number of electrons($2n^2$)
K – shell	1	$2 \times 1 = 2$
L – shell	2	$2 \times 2^2 = 8$
M – shell	3	$2 \times 3^2 = 18$
N – shell	4	$2 \times 4^2 = 32$

2. The outer most orbit of an element cannot contain more than 8 electrons and the orbit immediately before it can not contain more than 18 electrons .

3. It isn't always necessary to compete an orbit before the next orbit starts filling.

Ex- The electronic configuration of Sc is

$$_{21}Sc - 1s^2 2s^2 \, 2p^6 \, 3s^2 \, 3p^6 \, 4s^2 \, 3d^1$$
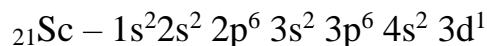
Here '4s' sub shell is filled before '3d' sub shell

Quantum Numbers :

Quantum number are the index numbers which specifies the address of the electron justifying it's position and energy in an atom.

There are 4 types of quantum number----

a) Principal quantum number (n)

b) Azimuthal quantum number (l)

c) Magnetic quantum number (m)

d) Spin quantum number (s)

Principle quantum number (n):

i)   It is designated as 'n' with values n = 1, 2, 3, 4, …..∞ indicating different orbits like K, L, M (shell).

ii)   Higher is the 'n' value greater is the size of the atom.

iii)   The energy of a given orbit is given by ----

$$E_n = -\frac{313.6}{n^2} Kcal / mol$$

iv) The maximum number of electron in any principal shell = $2n^2$ i.e the electrons in K, L, M, N shalls are 2 ,8, 18, 32 respectively.

b) **Azimuthal Quantum number** (l) :

i)   This is represented by symbol 'l' with values 0 to (n-1)

ii)   It signifies the presence of sub-shell in an orbit having different angular momentum eg.

For n = 1, the sub-shells are   l = 0 → s- subshell

For n=2,the sub-shells are   l= 0 → s- subshell

l=1→p-sub-shell

i.e the energy levels are further quantized

Hence, the jumping of electron from different sub-shells explains the fine lines.

iii)   The energy of different subshell in a principal shell follows in the order –

$s < p < d < f$

the maximum number of electron in a sub- shell $= 4l + 2$

for $l = 0 \rightarrow$ s - subshell $= 2$ electrons

        $l = 1 \rightarrow$ p - subshell $= 6$ electronss

        $l = 2 \rightarrow$ d - subshell $= 10$ electrons

        $l = 3 \rightarrow$ f - subshell $= 14$ electrons

iv) The shape of s,p,d and f are as follows-

sub-shell	shape
s	spherical
p	dumbell
d	double dumbell
f	complex

c)   **Magnetic quantum number (m)**

This quantum number is required to explain the splitting of spectral lines on application of magnetic field.

Thus the electron of a sub-shell can orient themselves in certain preferred regions of space around the nucleus called orbital's.

Magnetic quantum number determines the no. of required orientation of the electrons present in a sub shell

The no. of orientation in a sub-shell $= - l$ to $+l$

The total no. of orientation $= 2l + 1$

Sub-shell	values of "m"	no of orbitals
s(l = 0)	0	One
p (l – 1)	-1, 0, +1	Three
d (l = 2)	-2, -1, 0, +1, +2	Five
f(l=3)	-3, -2, -1, 0, +1, +2, +3	Seven

Thus, p- sub shell has three possible orientations i.e having three orbital's, p_x, p_y, p_z while d-sub shell have five orbital's such as dxy, dyz, dzx, dx^2-y^2, dz^2

d) **Spin Quantum number(s) :**

This helps to further quantized the energy levels as a result, fine spectrum can be explained.

It determine the spin of electron.

Spin Q. N has two values i,e +1/2 or -1/2

Thus electron may rotate in clockwise or anti- clockwise direction.

Pauli's Exclusion Principle :

No two electronics in an atom can have all the quantum number values alike.

Out of the four quantum numbers 3 may be identical but fourth quantum number value becomes excluded i.e non-identical. This principle can't be applied to hydrogen.

Ex –For He$\rightarrow 1s^2$

Electron	n	L	M	S
e_1	1	0	0	$+\dfrac{1}{2}$
e_2	1	0	0	$-\dfrac{1}{2}$

In the above example n, l, m quantum numbers values for two electrons one identical but the spin quantum number value differ.

Hund's rule of maximum multiplicity:

<u>Definition</u> :   Pairing of electrons doesn't take place in the degenerate orbitals of p, d and f sub-shell until each orbital contains a single electron with same spin.

The above is due to the reason that electrons being identical in charge repel each other when present in the same orbital. This repulsion can be minimized if two electrons move as far apart as possible by occupying different degenerate orbital.

Ex   Let us consider 3 electrons present in P subshell.

The   correct   way   of   arrangement   of   electron   in   orbitals   is Px^1 Py^1 Pz^1 (not Px^2 Py^1 Pz^0)

Aufbau's Principle:

Aufbau's is a German word with meaning building up. Thus, this principle helps in filling of the orbital with electrons. According to this principle the electrons are filled in different sub shell in order of increasing energy i.e the electron first occupies the lowest energy orbital and then enters to next higher energy orbital. The energy of different orbital's can be known from the $(n + l)$ rule.

Rule – 1 :

Lower is the (n+l) value lower is the energy.

For '1s' n=1 , l=0, n+l =1

For '2s' n=2, l=0, n+l=2

Thus, $E_{1s} < E_{2s}$

Rule -2 :

For sub-shells with identical (n+l) value ,lower is the energy for lower 'n' values.

For   4s,   n+l = 4 + 0 = 4

3p,   n+l = 3+1=4

Thus, $E_{3p} < E_{4s}$. as 'n' value for 3p is less.

Subshell	Energy(n+l)	Subshell	Energy(n+l)
1s	1+0= 1	4s	4+0=4
2s	2+0=2	4p	4+1=5
2p	2+1=3	4d	4+2=6
3s	3+0=3	5s	5+0=5
3p	3+1=4	5p	5+1=6
3d	3+2=5		so on

Thus increasing order of energy is…..

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

Electroni configuration of some element

$_1$H   1s1

$_2$He   1s2

$_3$Li   1s2 2s1

$_4$Be   1s2 2s2

$_5$B   1s2 2s2 2p1

$_6$C   1s2 2s2 2p2

$_7$N   1s2 2s2 2p3 (2px1 2py1 2pz1)

$_8$O   1s2 2s2 2p4 (2px2 2py1 2pz1)

$_9$F   1s2 2s2 2p5 (2px2 2py2 2pz1)

$_{10}$Ne   1s2 2s2 2p6 (2px2 2py2 2pz2)

Atomic Weight:

Atomic weight of an element is defined as the no. that an atom of an element is heavier than 1/12th mass of an atom of c^{12}

Thus,

Atomic mass $= \dfrac{Mass\ of\ an\ atom\ of\ any\ element}{\dfrac{1}{12}\ th\ mass\ of\ an\ atom\ of\ C^{12}}$

Average atomic mass:

The average atomic mass is calculated on the basis of isotopes and their abundance.

E.g. – Chlorine has two isotope Cl_{17}^{35} and Cl_{17}^{37} in the ratio 3:1.

So the average atomic mass of chlorine $= \dfrac{35 \times 3 + 37 \times 1}{3 + 1} = 35.5$

Molecular Weight :

Relative molecular mass of a substance is a number that shows how many times the mass of an atom is heavier than 1/12th mass of an atom of C^{12} isotopes.

Mol. Mass $= \dfrac{Mass\ of\ a\ molecule}{\dfrac{1}{12}\ th\ mass\ of\ an\ atom\ of\ C^{12}}$

Equivalent Weight :

Equivalent mass of a substance is the number of parts by mass that can combine or displaces 1.008 parts by mass of hydrogen, 8 parts by mass of oxygen or 35.5 parts by mass of chlorine.

The equivalent mass expressed in gm is called gm equivalent mass.

Relationship between atomic mass, equivalent mass and valency-

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Atomic mass = Equivalent mass x valency

CHEMICAL BONDING

Bonding is the combination of atoms to form a molecule. Atoms combine to have stable electron configuration or to satisfy octet rule.

Definition: A chemical bond is defined as a force of attraction which holds together the constituent atoms in a molecule.

Types of chemical bond :

There are 3 main types of chemical bonds. Such as :

a) Electrovalent / Ionic bond

b) Covalent bond

c) Dative bond or co-ordinate bond

Ionic bond / Electrovalent bond :

Features:

a) Here the bond is formed between two dissimilar atoms (A & B).

b) Atom 'A' has tendency to lose one or more electron by occupying stable electronic configuration and form (+) vely charged ion (cation)

Eg – $Na(g) - \bar{e} = Na^+(g)$ (Iso electronic with Ne)

c) Another atom 'B' has tendency to accept one or more electron occupying stable electronic contiguration and formed (-) vely changed ion (anion).

$Cl(g) + e^- = cl^- (g)$ (Iso electronic with Ar)

d) These ions are held by strong electrostatic force of attraction as a result the potential energy is decreased.

e) The electrovalence of an atom is the no. of electron that it can gain or lose to have stable electronic configuration.

Factors affecting electrovalent bond formation:

a) Ionization Energy (I.E) :

Lower is the I.E easier is the formation of cation and stronger is the ionic bond.

b)   Electron affinity :

Higher is the electron affinity easier is the formation of anion and stronger is the ionic bond.

c)   Force of attraction :

Stronger is the force of attraction between the oppositely charged ions stronger is the ionic bond.

Explanation of the formation of Ionic bond.

Example of the formation of $CaCl_2$

The electronic configuration of Ca and Cl are given below, such as -

Ca_{20} $\rightarrow$ 2   8   8   2

Cl_{17} $\rightarrow$ 2   8   7

Ca^{2+} $\rightarrow$ 2   8   8

Cl^- $\rightarrow$ 2   8   8

$Ca^{2+} + 2Cl^- \rightarrow CaCl_2$

The electrostatic force of attraction between Calcium and chloride ions result in the formation of $CaCl_2$.

Ex – KCl, KBr, KI, CaF_2 , $CaBr_2$

Covalent Bond :

Features-

(i)   A covalent bond is formed between two similar or dissimilar atoms (A – A or A-B)

e.g – (H – H, H-Cl)

(ii)   Here both the atoms are short of electrons to have stable configuration.

(iii)   Both the atoms share mutually such that each atom occupies stable configuration.

(iv)   According to covalent bond molecules have two types of electrons------------

a)Bond pair of electrons / shared pair of electrons i.e those take part in smutual sharing .

b)Unshared pair of electron/ lone pair of electron.

These pair of electrons doesn't take part in sharing but satisfies octet.

Ex- In Oxygen molecule each oxygen atom has two electrons as bond pair, where as other 4 electrons present as 2 lone pairs

Fig---

LP — oxygen diagram — LP

sDative bond/ Co-ordinate bond :

Features:

 i) Like covalent bond this bond is formed between two similar or dissimilar atoms.

ii)Here one of the atom is short of electrons two have duplet or octet configuration and such atoms are called acceptor atom.

 iii) Another atom is having an excess pair of electrons(lone pair) and termed as donor atom.

 iv)   The donor atom partially shared a pair of electron with the acceptor atom such that both the atoms satisfy stable electronic configuration (duplet / octet)

iv)   The bond is represented by an arrow heading towards the acceptor atom from the donor atom.

Fig

— A ⟶ B   ≡   — A ⟶ B
Acceptor Atom
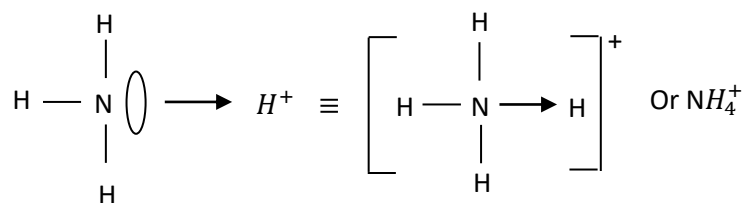Donor Atom
Co-ordinate bond
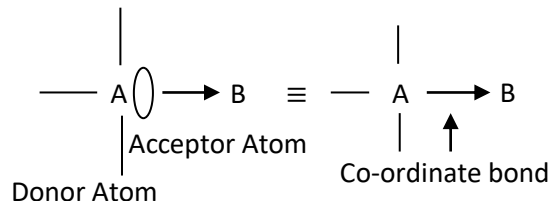
vi)   Thus the chemical bond formed by partial donation and partial sharing of lone pair of electron between two atoms or ions is called a co-ordinate / dative bond.

Ex – NH_4^+ , H_3O^+

Formation of ammonium ion-

Fig

$$H \longrightarrow N \xrightarrow{\quad} H^+ \equiv \left[H \longrightarrow N \longrightarrow H \right]^+ \quad \text{Or N}H_4^+$$
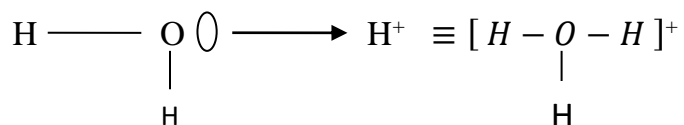
.

Here one atom can share the electron pair but both will satisfy the stable configuration .

'N' satisfies octet and 'H' satisfies the duplet state.

Formation of hydronium ion -----

Figs

$$H \longrightarrow O \longrightarrow H^+ \equiv [\, H - O - H \,]^+$$

Here Oxygen satisfies Octet and 'H' satisfies duplet.

1.2 Arrhenius theory of Acid Base:

According to this <u>Acids</u> are the substances that can provide H^+ ion in aqueous medium

ex : HCl, HNO_3, CH_3COOH etc

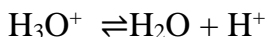<u>Bases</u> are those that can provide $\overline{OH}$ ion in aqueous medium.

Ex: NaOH, $Ca(OH)_2$ $Al(OH)_3$

Bronsted – Lowry Theory :

According to this theory acids are those substances capable of donating proton or acids are proton donors.

Base are those species capable of accepting proton or bases are proton acceptor.
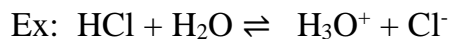
Example of Acid ;

$$H_3O^+ \rightleftharpoons H_2O + H^+$$

$$NH_4^+ \rightleftharpoons NH_3 + H^+$$

Example of Bases :

All anions are Bronsted Lowery base.

Ex: Cl^-, SO_4^{2-}, PO_4^{3-} etc.

By proton transfer conjugate pair of acids and bases are formed.

Ex: $HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$

　　(acid) (base)　　(acid)　(Base)

Hence Cl^- is a conjugate base of acid HCl

H_3O^+ is a conjugate acid of base H_2O.

Thus,　acid $- H^+$ = conjugate base,

Base $+ H^+$ = conjugate acid

$$NH_2^- \xleftarrow{\;-H^+\;} NH_3 \xrightarrow{\;H^+\;} NH_4^+$$

(Conjugate base)　　　　(conjugate acid)

Lewis Theory :

According to Lewis theory

- An acid is a substance which can accept a pair of electrons from any other substance i.e electrons acceptors are acids.

- A base is a substance which can donate a pair of electrons to any other substance. i.e electron pair donors are bases.

Examples of acids :

a) All cations are Lewis acids. Such as Na^+, K^+, Ca^2, Cu^{2+}, Al^{3+} Fe^{3+} etc.

b) Neutral molecules containing electron deficient atoms are Lewis acids

Eg – BF_3, $AlCl_3$, $feCl_2$, $ZnCl_2$ etc

Examples of Bases :

- All anions are Lewis bases – F^-, Cl^-, Br^-, OH^-, CN^- NO^-_2, O^{2-}, S^{2-}

Salt :

It is formed by complete neutralisation of aqueous solution of strong acid with an aqueous solution of a strong base.

Neutralization of Acid and Base :

Salts are formed out of acid base reactions by replacement of H^+ ion of an acid by any metal ion or other cation. Thus, these are electrovalent compound of which the cations come from a base (basic radical) and anions come from an acid (acid radical)

Ex:

$$NaOH + HCl \rightarrow NaCl + H_2O$$
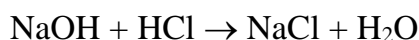
Here, Na^+ is a part of base NaOH and Cl^- is a part of acid HCl.

Classification of Salt : (On action of water)

a) Salts of strong acid and strong base :

NaCl / KNO_3, the solution is neutral.

b) Salts of strong acid and weak base

Ex – NH_4Cl - the solution is acidic

$$NH_4Cl + H_2O \rightarrow NH_4OH + HCl \text{ (strong acid)}$$

c) Salt of strong base and weak acid :

Ex- CH_3COONa the solution is alkaline

$P^H > 7$

d) Salt of weak acid and weak base :

Eg – CH_3COONH_4

These type of salts are either acidic or alkaline depending on the degree of dissociation of acid and base / relative proportion of H^+ and OH^- ions in solution. In case, the acid is relatively stronger compared to the base, the solution will be acidic. Also a relatively stronger base in the neutralization reaction with a weak acid forms a basic solution.

Equivalent weight of acid, base & salt :

Equivalent weight of Acid = Mol mass/ Basicity

Where Basicity = No of H^+ ion replaceable

Ex- Equivalent mass of $\quad HCl = \dfrac{36.5}{1}$

$$H_2SO_4 = \dfrac{98}{2} = 49$$

Equivalent weight of base $= \dfrac{Mol.mass}{Acidity}$

Acidity = No. of $\overline{OH}$ ion replaced.

Eg. – Equivalent mass of NaOH = 40

Equivalent mass of $Ca(OH)_2 = \dfrac{74}{2} = 37$

Equivalent mass of $Al(OH)_3 = \dfrac{Mol.mass}{3} = \dfrac{78}{3} = 26$

Equivalent weight of salt $= \dfrac{Mol.mass}{Total\,charge\,on\,cation\,/\,anion}$

Eg :   Equivalent mass of $Na_2SO_4 = \dfrac{Mol.mass}{2}$

Equivalent mass of $(Al)_2\,(SO_4)_3 = \dfrac{Mol.mass}{6}$

Standard Solution :

The solution whose strength is known is called standard solution.

Normal Solution ;

  when one gram equivalent of a Solute is dissolved in one liter solution, then it is termed as Normal Solution

Ex- Equivalent weight of NaOH =40 .

When 40 gm of NaOH dissolved in 1 lit solution, normal solution of NaOH is formed.

Molar solution:

When one mole of a solute is dissolved in one litre solution, it is called a molar solution.

Ex –Molecular mass of $NaCO_2$ =106

When 106gm of Na_2CO_3 is dissolved in 1 lit solution, one molar solution of Na_2CO_3 is formed.

Molal Solution :

When 1 mole of solute present in 1000gm of a solvent, it is called 1 molal solution.

Ex: 106 gm of Na_2CO3 in 1000gm of solvent.

Normality – (N) :

It is defined as the no. of gram equivalent of the solute present per litre of the solution at a given temperature.

$$N = \frac{no.of\ gram\ equivalents}{Vol^m\ of\ solution}$$

$$N = \frac{\omega}{\in} \times \frac{1000}{V}$$

$\frac{\omega}{\in}$ = No. of gm equivalent.

W = wt. of the solute

E= Equivalent wt. of the solute.

If 1 gm equivalent present in 1 litre solution i.e called 1 N solution.

Molarity (M) It is the no. of moles that present per liter of solution at a given temperature.

$$Molarity(M) = \frac{\omega}{M} \times \frac{1000}{V}$$

Where w = wt of the solute

M= molecular mass of solute

v= volume of solution

Molality(m) :

It is the no. of mole that present per kg of solvent.

Molality (m) = $\dfrac{\omega}{M} \times \dfrac{1000}{\omega'}$

Where → ω = wt. of the solute ω'= wt. of the solvent.

M = Mol. Mass of the solute

1Molal solution means 1 mole of solute present in 1 kg of water.

pH :

pH of the solution is the power of hydrogen ion concentration in the solution .

Mathematically : It is expressed as (-) ve logarithm of molar concentration of hydrogen ion

i.e pH= - log [H⁺]

or   pH = - log $\dfrac{1}{[H^+]}$

so, pH decreases with rise in concentration of hydrogen ion.

pH scale ranges from O-14 with a midpoint at '7'

Importance of PH value in industries

1.   **Water for domestic use :**

The pH values of the public water supplied should preferably be between 7.2 to 8.0. If pH < 7.2 then corrosion may occurs and pH > 7 indicate certain physiological effects.

2.   **Water treatment**:

pH control is required for treatment of sewage waste by coagulation process, other wise the added coagulants are not fully hydrolysed and they are in effective to bring about sedimentation of fine suspended and colloidal impurities. Optimum pH value various from 4.2 to 7.6 , when alum is used as cation co-agulant.

3.   **In Sewage Treatment**:

The pH value from 7.3 to 7.56 is required to decide which particular organism will bring about the decomposition most satisfactorily.

4. **Corrosion** : In pH < 5, the corrosion of iron structure in oxygen free water is slow, but the rate of corrosion is very high in presence of oxygen at pH 4 to 5 and 10 to 12.

5. **In chemical industry :**

pH adjustment is required to have a proper control over the reaction.

S6.   Sugar Industry :

The pH of the juice is maintained to 7. If it is acidic (i.e pH< 7) the sucrose in the juice is hydrolyzed to glucose and fructose mixture. But if the juice becomes alkaline (pH > 7) undesirable acids and coloured substances are produced.

7.   In electroplating :

To have adherent, smooth and shinning deposits, the pH of the electroplating both is to be carefully controlled.

Buffer Solution :

A buffer solution is defined as a solution which maintained the change in pH value, when a little quantity of acid, base or water is added into it.
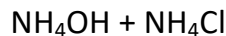
Types of Buffer Solution :

i)   **Acidic Buffer :** Buffer solution having pH less than 7 .

Ex – Weak acid with it's salt.

$CH3COOH + CH_3COONa$

ii)   **Alkaline Buffer :** Buffer solution with pH greater than 7.

Ex – weak base with its salt.

$NH_4OH + NH_4Cl$

Thus a buffer contains a weak acid and it's conjugate base or weak base and it's conjugate acid. These are called buffer pairs.