

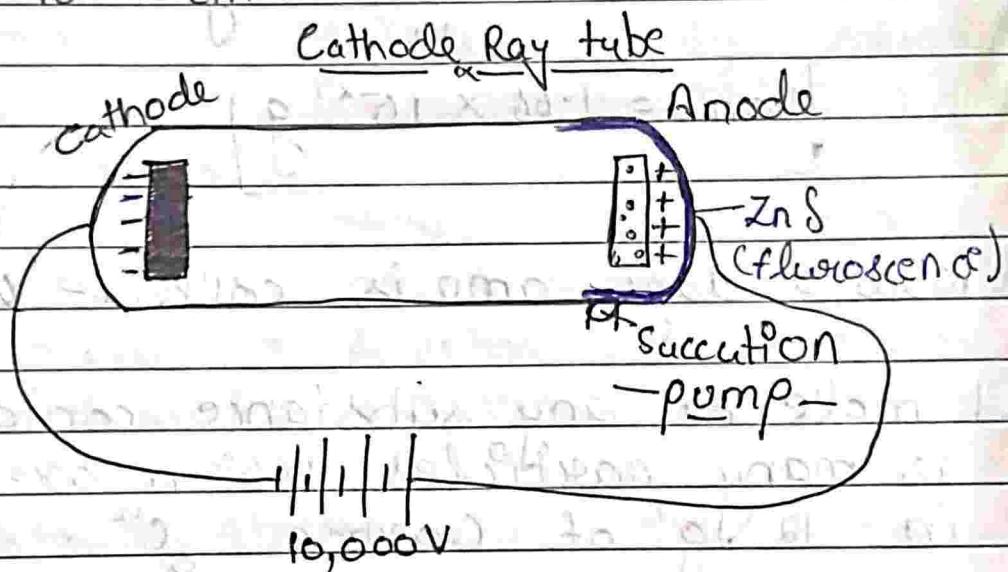
# CHAPTER:- Atoms

## # Discovery of Electrons :-

[JJ thompson's Cathode Ray from Tube Ex.]

Conditions :-

- ① Low Pressure
- ② Inert gases
- ③ High Potential (10,000 v)
- ④ Metallic Electrodes (Aluminium)
- (Because they are easily ionisable)
- ⑤ Gap at 10-20 cm



## # Properties of Cathode Rays

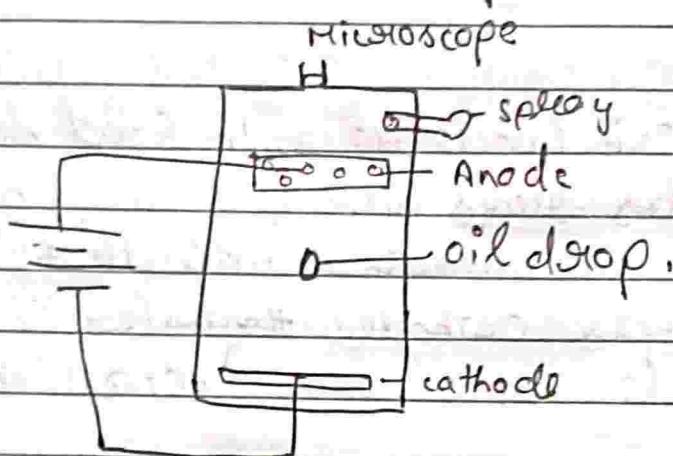
- ① The rays coming from cathode is called cathode R.
- ② These rays are moving from cathode to Anode
- ③ Cathode rays always travels in straight line ( $c = 3 \times 10^8$ )
- ④ Cathode rays contains particle which have mass.
- ⑤ Cathode rays have momentum ( $p = mv$ ).
- ⑥ Particles of cathode rays are negatively charged and called electrons ( $e^-$ ).

# Specific Charge (charge by mass Ratio)  
 " J thomson experimentally found that,

$$e/m = 1.75880 \times 10^{-11} \text{ C/kg}$$

\*  $e/m$  for any gas will remain same.

# "Millikan Oil Drop Experiment":



By this experiment Millikan found the charge of  $e^-$ : that is,

$$\boxed{e = -1.67 \times 10^{-19} \text{ C}}$$

this is a integral value.

"with this and Specific charge Ratio of J.J. thomson: mass of  $e^-$  is calculated"

$$\boxed{m = 9.1 \times 10^{-31} \text{ kg}}.$$

# Discovery of Proton :-

(E Goldstein's Canal Ray Experiment)

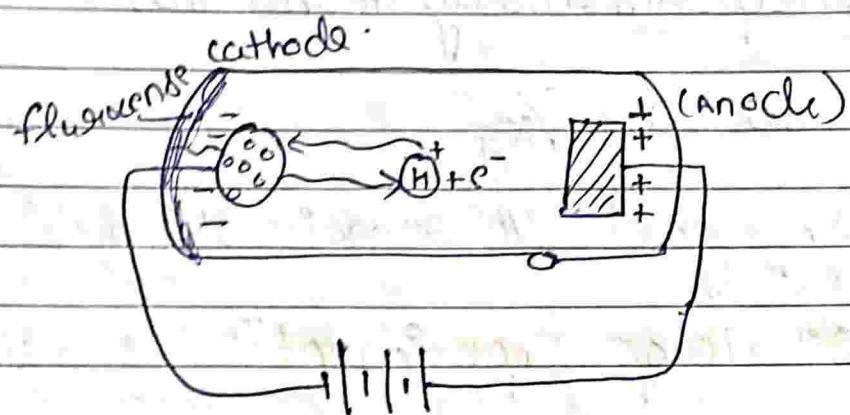
Conditions :-

① Low pressure

②  $H_2$  gas

③ High potential

## (Q) Metallic electrodes



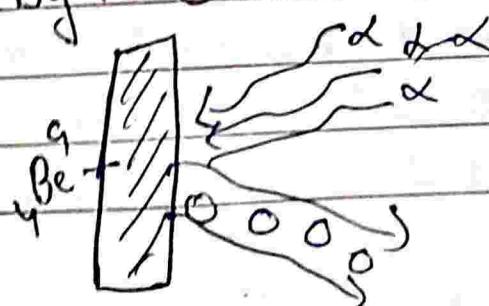
# (There is no thing like Anode rays, In Goldstein's experiment cathode rays hits H<sub>2</sub> atom with very high speed and breaks it into H<sup>+</sup> + e<sup>-</sup>. H<sup>+</sup> moves towards cathode & transfer through pores of cathode so, known as canal rays.)

### Properties:

- ① The rays are canal rays.
- ② The particle they contain is positively charged called proton.
- ③ Charge on proton  $p^+ = +1.67 \times 10^{-19} C$ .
- ④ Mass of proton  $m = 1.672 \times 10^{-27} kg$

$e/m$  here will depend on gas.

# Discovery of Neutron  
(By : Chadwick)



- \* Particle emitted from Be plate were neutral.
- (a) called neutrons. ~~without~~
- \* Mass of neutron  $m = 1.674 \times 10^{-27} \text{ kg}$ .
- \* These are heaviest particles.

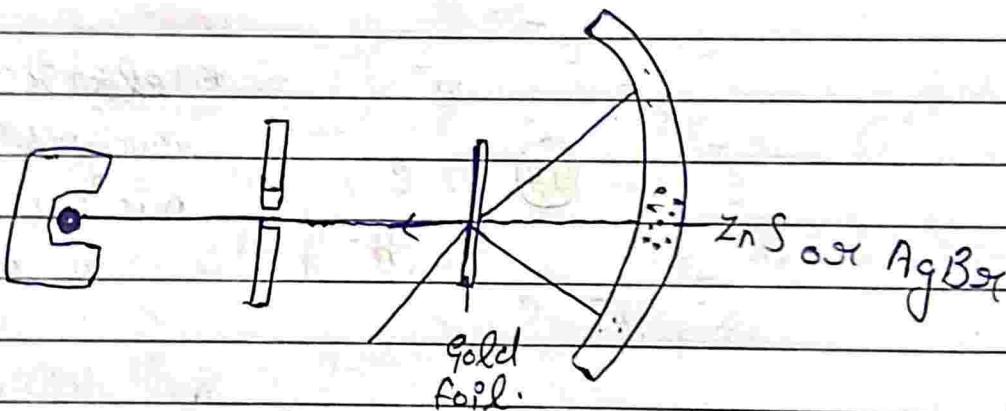
$$n^0 > p^+ > e^-$$

## # Rutherford $\alpha$ -Scattering Experiment:-

Radioactive Substances emits :-

$$\begin{aligned} \alpha\text{-particles} &\rightarrow {}_{+2}^{4\text{amu}} \alpha \\ * \text{Massive particles} & \quad \left. \begin{aligned} 1\alpha &= 2p^+ \text{ (charge)} \\ 1\alpha &= 4p^+ \text{ (mass)} \end{aligned} \right. \end{aligned}$$

$({}_{2}^{4}\text{He} \text{ nuclie})$



Observation :-

① Most of the  $\alpha$  particles passed through the gold foil in a straight line.

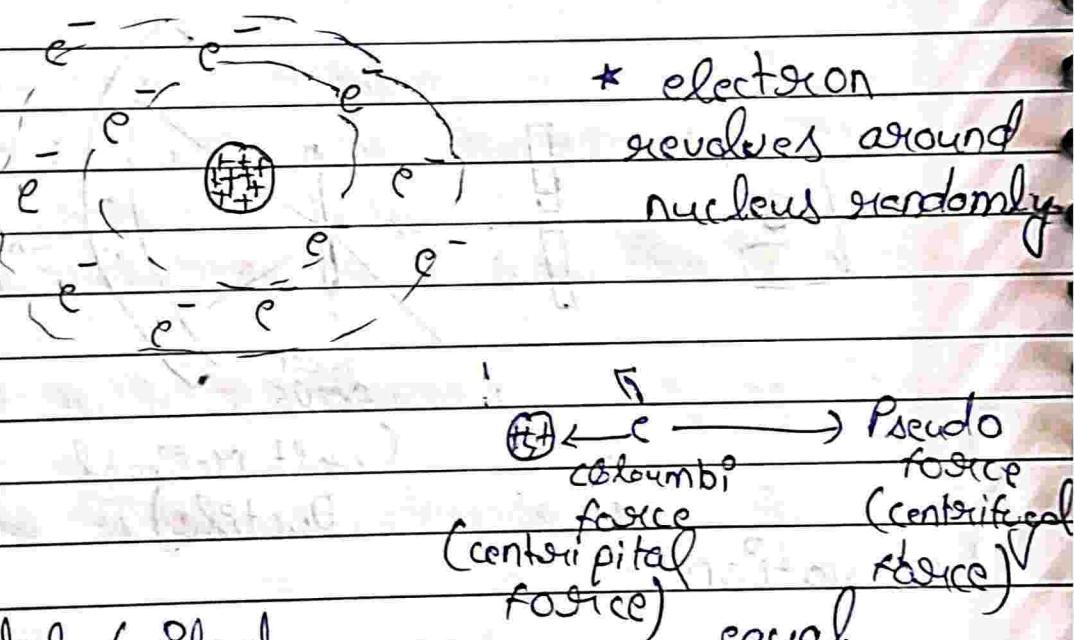
② Out of every 200-particles 1  $\alpha$ -particle deviates its path.

③ Out of every 20,000 particles 1 particle deviated at  $180^\circ$ .

### Conclusions:

- ① Most of the part of atom is empty.
- ② At the centre of the atom all the positive charge (Protons) are accumulated.
- ③ This central have all the mass of the atom. this called nucleus.

### Rutherford Model of Atom



$\text{(+)} \leftarrow e \rightarrow$  Pseudo force  
Coulombic force (centrifugal force)

~~This model failed~~ with in 15 mins of acceptance. (Conflicted by Neil Bohr in reference to James Clark Maxwell's law)

opposite force

## # Some Important terminologies :-

1. Atomic Number (Z) :- No. of  $e^-$  or  $p^+$  present in an atom, in case of ions no. of  $p^+$  present.

2. Atomic Number (A) :- Sum of number of protons and neutrons in an atom.

3. Representation :-  $\text{X}^A$

4. Isotopes :- Elements have same Z but different A. Ex:-  ${}_{6}^{12}\text{C}$ ,  ${}_{6}^{13}\text{C}$ ,  ${}_{6}^{14}\text{C}$

5. Isobars :- Elements that have same A but different Z. Ex:-  ${}_{19}^{40}\text{K}$ ,  ${}_{20}^{40}\text{Ca}$ ,  ${}_{18}^{40}\text{Ar}$

6. Isoelectronic :- Elements or ions that have same number of electrons. Ex:-  $\text{O}^{2-}$ ,  $\text{F}^-$

## Limitation of Rutherford Model of Atom

1. According to James Clark Maxwell's Rules charges in random motion, the electron will collides to a nucleus and destroys the atom but it doesn't happen.

2. There is no fixed position or placement of electron.

3. According to this model atom should give continuous spectrum but it gives line spectrum.

# Development before Neil Bohr's Model :-

① Electromagnetic Waves/Radiations [EMR] :-  
"Energy carriers, Invisible/visible, travels with speed in a straight line, can travel in vacuum, dual nature, they create  $E_x$  and  $B_z$  but they nullify, does not effected by MF and EF."

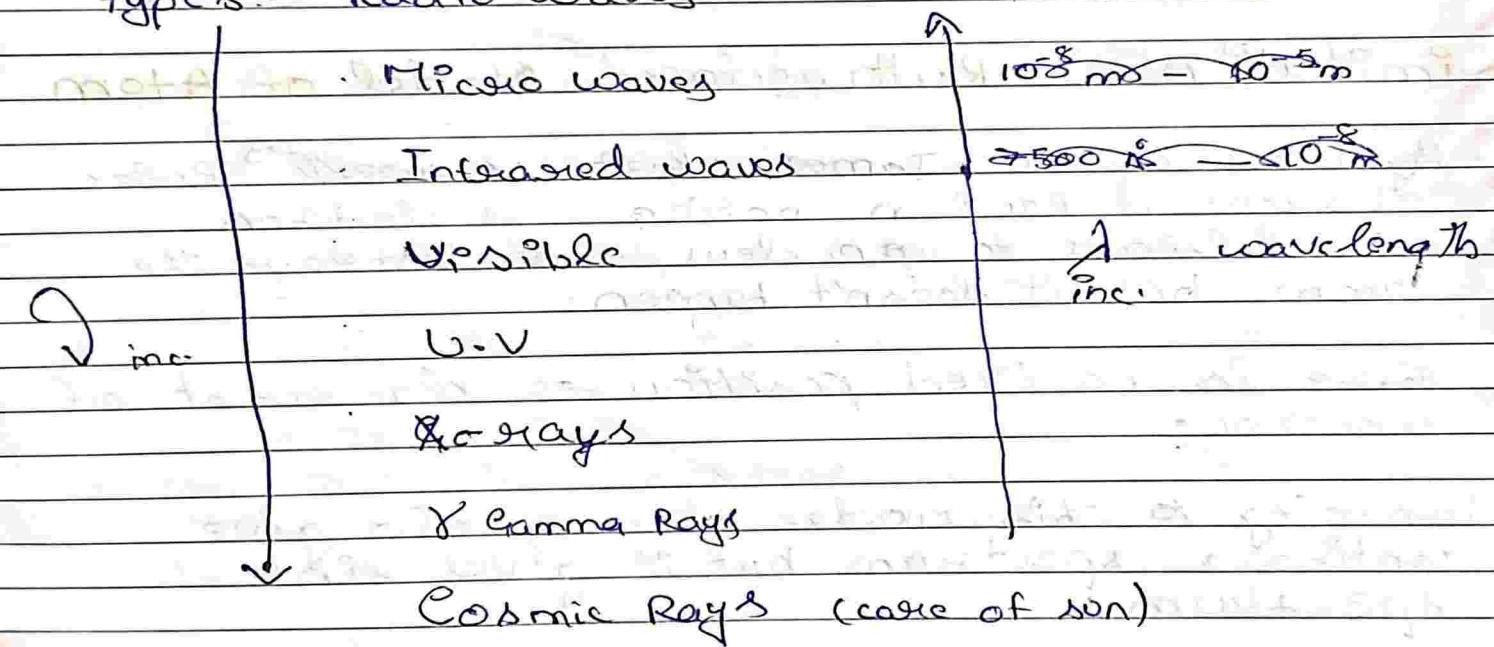
\* wavelength :- length of one wave. unit - m (1)

\* frequency :- no. of waves passing through a point in (v) one sec. unit:- Hertz,  $s^{-1}$

$$\star \quad \checkmark = \frac{c}{1}$$

$$\star \text{ wave number } (\overline{\lambda}) = \frac{1}{\lambda} \quad \begin{matrix} \text{No. of wavelength} \\ \text{per unit length.} \\ \text{unit} = \text{m}^{-1} \end{matrix}$$

types:- Radio waves



## # Plank's Quantum Theory

"EMW travels in the form of Energy packets, Energy bundles called Photons"

$$E \propto V \Rightarrow [E = hV] \quad h = \text{Plank's constant}$$

### \* Black body :-

"Ideal body that emits and absorbs all the Radiations"

There is no black body on earth, Sun can be considered as Ideal black body.

### # Spectrum :-

"splitting of light"

These are of Three types :-

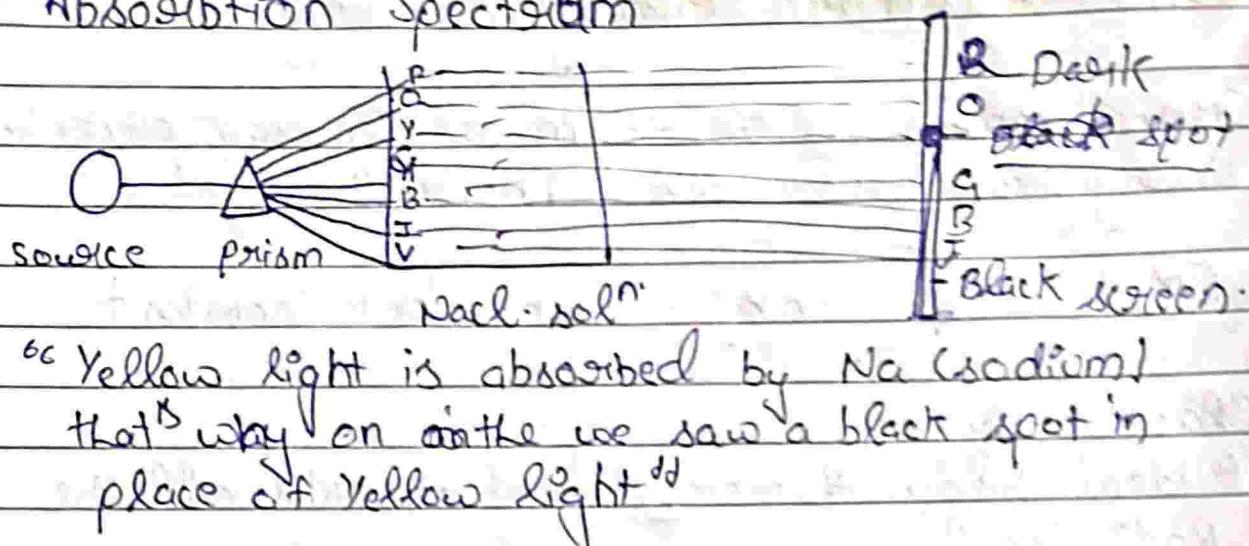
#### ① Emission Spectrum

(a) Continuous Spectrum :- "all types of EMR can be observed" only sun can form continuous spectrum.

(b) Band Spectrum :- "A specific EMW and its complete Range is observed"  
Ex:- White light shows VIBGYOR.

(c) Line Spectrum :- A spectrum which provide the specific EMW of specific Radiations.

## ② Absorption Spectrum

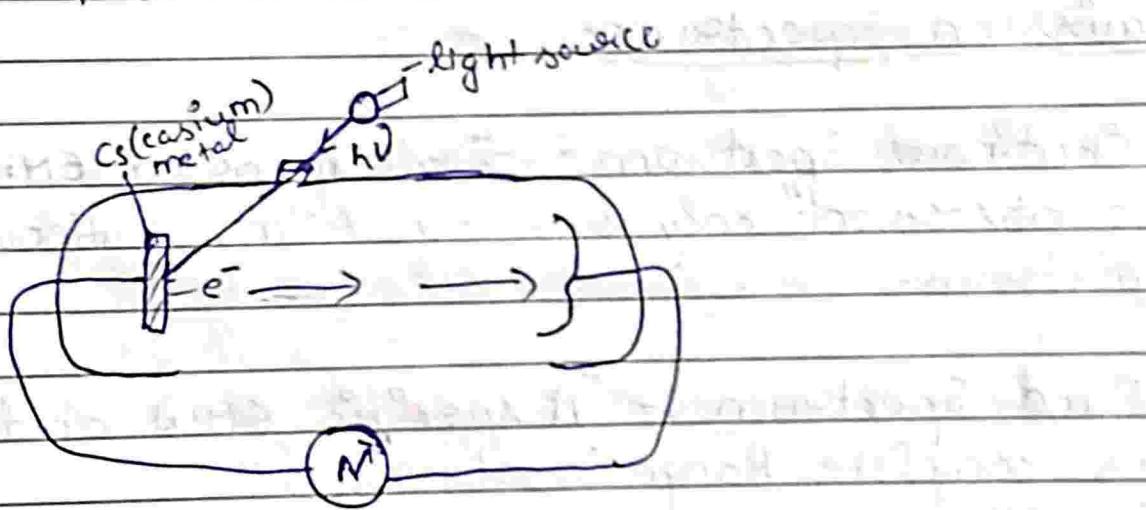


"Yellow light is absorbed by Na (sodium) that's why on the we saw a black spot in place of yellow light"

"When a specific radiation's Energy is absorbed and other radiations in EMW are visible that spectrum is called absorption spectrum"

## # Lenard's and Hertz Experiment

"Photoelectric effect"



Conclusions:-

- ① Photoelectric emission is instantaneous process  
"There is no time lag b/w striking of photon and emission of electron"

## ② Photocurrent & Intensity of Radiation.

As Intensity increases ↑  
 ↓

Number of Photons striking increases ↑  
 ↓

Number of Electrons emitted increases ↑  
 ↓

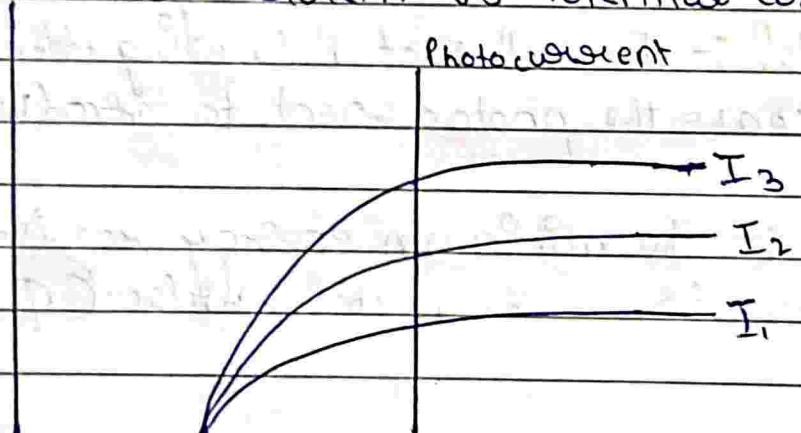
Photocurrent increases.

## ③ Value of frequency before which there is no photocurrent

"The minimum frequency needed to emit electrons is called Threshold frequency"

## ④ K.E (kinetic energy) of electrons depends on frequency of Radiations not intensity of Radiation

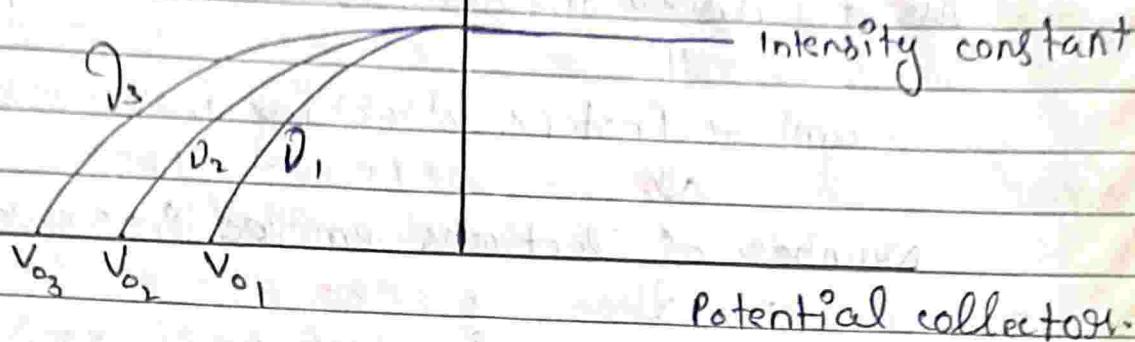
Photocurrent vs Potential collector (diff. I same)



$V_0$  (stopping Potential)

$$I_3 > I_2 > I_1 \quad I = \text{constant}$$

## Photo current



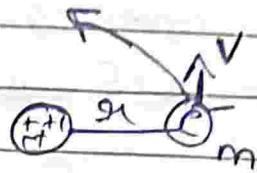
$$J_3 > J_2 > J_1$$

- ① Threshold Frequency :- The minimum frequency of a Photon required to emit an electron from the surface of metal. ( $\nu_0$ )
- ② Threshold Wavelength :- The maximum wavelength of a photon required to emit an electron from the surface of metal ( $\lambda_0$ )
- ③ Stopping Potential :- The value of Retarding Potential required to cease the photocurrent to zero ( $V_0$ )
- ④ Work function :- The minimum energy required to eject electron from the metal surface ( $\phi$ )

## # Neil Bohr's Model of Atom :-

1. Atom contains nuclei & it passes all the n & p<sup>+</sup> in it, it is very massive.

2.  $\left\{ \begin{array}{l} \text{Angular momentum} \\ = mv r_l \end{array} \right\}$



3. Electrons do not revolves randomly, e<sup>-</sup> always move a specific path called orbit at specific radius. [Shell  $\rightarrow K, L, M, N$ ]. with specific total energy V<sub>T</sub>

4. When e<sup>-</sup> absorbs energy it gets excited and reaches higher energy levels orbital.

5. When the excited e<sup>-</sup> releases all its extra energy & and reaches to its ground state.

6. The angular momentum of an e<sup>-</sup> in a shell is always equal to  $\frac{nh}{2\pi}$ , n = no. of shell, h = planck's constant

\*  $\frac{nh}{2\pi} \rightarrow$  Bohr's Quantization.

\* Radius of Shell

$$r_l = \frac{n^2 h^2}{4\pi^2 m e^2 k Z} = 0.529 \text{ Å} \frac{n^2}{Z}$$

\* Velocity of e<sup>-</sup>

$$v = \frac{nh}{2\pi m r_l} = \frac{e^2}{2\pi n \epsilon_0} = 2.18 \times 10^6 \times \frac{Z}{n}$$

\* Total Energy of a shell =  $-\frac{13.6}{n^2} Z^2 \text{ eV}$

also T.E =  $-\frac{ke^2}{2r} \cdot Z$ , P.E =  $-\frac{ke^2}{r}$ , K.E =  $\frac{ke^2}{2r}$

\*  $\Delta E = E_2 - E_1$ ,  $\Delta E = h\nu = \frac{hc}{\lambda}$

$$\nu = 1.097 \times 10^7 \left( \frac{1}{n_f^2} - \frac{1}{n_i^2} \right) Z^2$$

$$R_H = 1.097 \times 10^7 \text{ (Rydberg's Constant)}$$

If.  $n_f = 1 \Rightarrow$  Lyman Series (U.V)

$n_f = 2 \Rightarrow$  Balmer Series Visible

$n_f = 3 \Rightarrow$  Paschen Series Infrared

$n_f = 4 \Rightarrow$  Brackett Series -

$n_f = 5 \Rightarrow$  Pfund Series (Visible range)

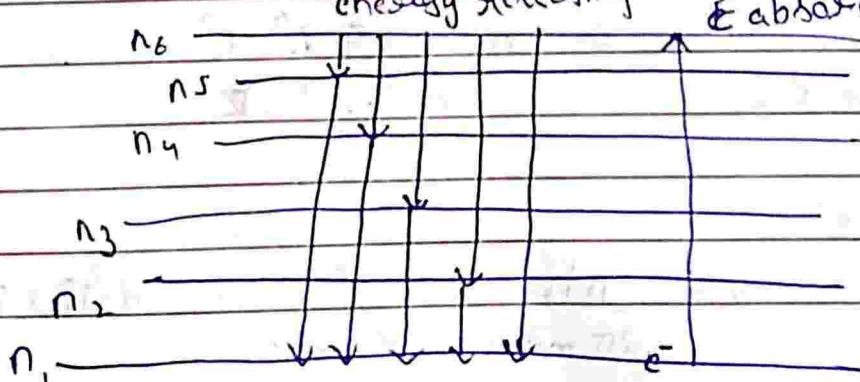
$$\Delta E = 13.6 \left( \frac{1}{n_f^2} - \frac{1}{n_i^2} \right) Z^2$$

no. of possible transitions or no. of spectral lines released

$$= \frac{(n_f - n_i)(n_f - n_i + 1)}{2}$$

deexcitation  
energy releasing

Excitation  
 $E_{\text{absorbed}}$



## # De-Broglie Dual Nature of matter

According to De-Broglie dual nature of matter during the revolution of  $e^-$  around nucleus in an atom they continuously emit some waves called De-Broglie waves

"Something which have mass and momentum will also have wave nature"

$$\text{E} = \frac{hc}{\lambda} \quad \text{--- (1)}, \quad E = mc^2 \quad \text{--- (2)}$$

$$\frac{hc}{\lambda} = mc^2$$

$$\lambda = \frac{h}{mc} \quad \text{--- (3)} \Rightarrow \boxed{\lambda = \frac{h}{p}}$$

also by eq (3)  $K.E = \frac{1}{2}mv^2$

$$2m \times KE = m^2 v^2$$

$$\boxed{KE = \frac{p^2}{2m}} \quad \text{--- (4)}$$

from eq (4)

$$p^2 = K.E \times 2m$$

$$p = \sqrt{2mKE}$$

$$\boxed{\lambda = \frac{h}{\sqrt{2mKE}}}$$

It also confirms Bohr's quantization as :

$$\lambda = \frac{h}{mv}$$

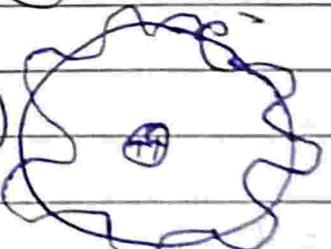
$$mv = \frac{h}{\lambda} \quad \text{--- (1)}$$

$$\text{also, } 2\pi r = n\lambda \quad \text{--- (2)}$$

(as  $e^-$  completes a circle and covers  $n$  no. of waves)

$$mv = \frac{h}{\frac{2\pi r}{n}}$$

$$mvn = \frac{nh}{2\pi} \quad \text{angular momentum of } e^- \text{ by Bohr.}$$

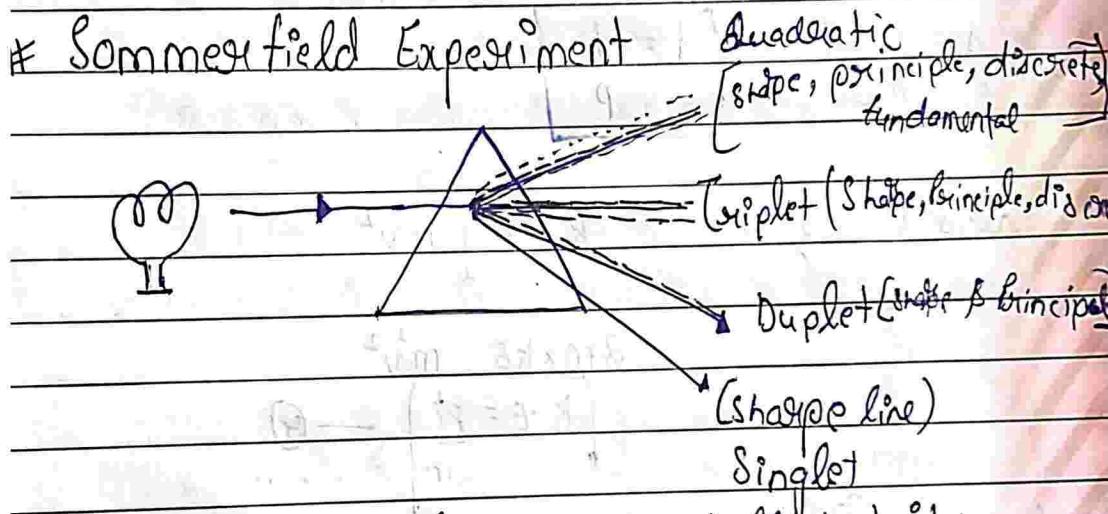


## # Heisenberg Uncertainty Principle

When we try to obs. the exact position of an  $e^-$ ; its velocity will change and if we want to know its velocity than its position is uncertain.

$$\Delta x \cdot \Delta p \geq \frac{h}{4\pi}, h = \text{plank's constant.}$$

Hence, An atom do not have fixe / accurate / certain trajectories but they have some irregular path.



\* Atom do not only contains shell but its also contains subshells.

Shell	Subshell	Name
$n=1$	1.	S - subshell (sharp)
$n=2$	2	S, P - subshell (Principle)
$n=3$	3	S, P, D - subshell (discrete)
$n=4$	4	S, P, D, F - subshell (fundame)

## # Limitations of Neil Bohr's Atom

1. This model was only applicable for Hydrogen & hydrogen like atom.
2. Sommerfeld Explains presence of subshells.
3. Zeeman Effect :- splitting of spectra on electromagnetic waves due to high magnetic field because they have subshells.
4. Stark Effect :- splitting of spectra due to high Electric waves.
5. It does not tell anything about chemical bond and molecule formation.

## # Latest Quantum Mechanical Model of Atom.

By :- Schrodinger.

$$\frac{d^2\psi}{dx^2} + \frac{d^2\psi}{dy^2} + \frac{d^2\psi}{dz^2} + \frac{8\pi^2 m(E-V)}{\hbar^2}\psi = 0$$

$\psi$  = wave function , it does not have any physical significance.

$\psi^2$  = probability of finding  $90^\circ$  of  $e^-$   
 $x, z$  ↑  
 $x, y, z \rightarrow$  axis       $\neq y$

$V$  = potential Energy ,  $E$  = Energy ,  $m$  = mass ,  $\hbar$  = planck constant

\* Hence the path of electron is not that specified  
 So, it is not possible to find radius and  
 velocity of Bohr acc. to Bohr. But we  
 can find its energy by  $E = -\frac{13.6 Z^2}{n^2} \text{ eV}$ .

## # Quantum Numbers.

1. Principle quantum numbers (n)
2. Azimuthal quantum numbers (l)
3. Magnetic Quantum numbers (ml)
4. Spin quantum numbers (s)

### ① Principle quantum number :- (n)

"e<sup>-</sup> is situated in which shell/orbit"

Energy; Size.

$$n=1 < n=2 < n=3 < n=4$$

$n=1$  (first shell) K

$n=2$  (Second shell) L

$n=3$  (third shell) M

$n=4$  (fourth shell) N

### ② Azimuthal quantum Number :- (l)

"it tell us about subshell in which e<sup>-</sup> is present"

Representation:- 0 to l

$$\boxed{l=n-1}$$

Example:- if  $l=1$

$$l=0, 1$$

$l=0$  (s-subshell)

$l=1$  (p-subshell)

$l=2$  (d-subshell)

$l=3$  (f-subshell)

③ Magnetic quantum Number :-  $m_l$

"e<sup>-</sup> cloud / an area in atom where probability of find e<sup>-</sup> are 90°"

⇒ In one orbital = 2e<sup>-</sup> can be accommodated

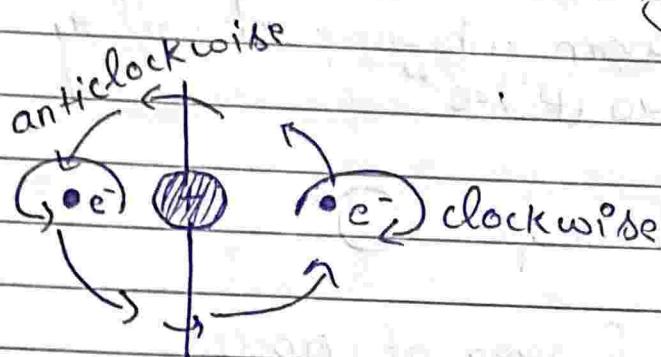
$$m_l = 2l + 1$$

Representation =

$$m_l = -l \text{ to } l$$

④ Spin Quantum Number.

"Inside a single orbital"



# opp. spin both e<sup>-</sup> will cancel each

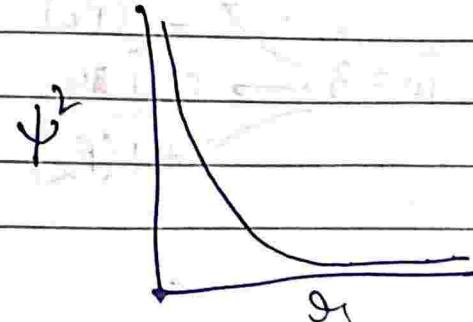
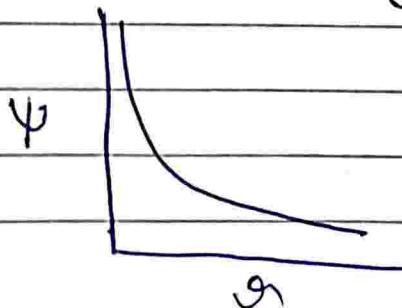
other magnetic moment.

represented as  $+\frac{1}{2}$  &  $-\frac{1}{2}$

Shell → Subshell → Orbitals → electron  
(Orbit) (Energy) 1 or L  
level

# Shape of Orbitals (S)

$\Psi$  &  $\Psi^2$  vs  $r$  graph for 1s orbital.



$\rightarrow$  S-graph  
upper saishwar hoga

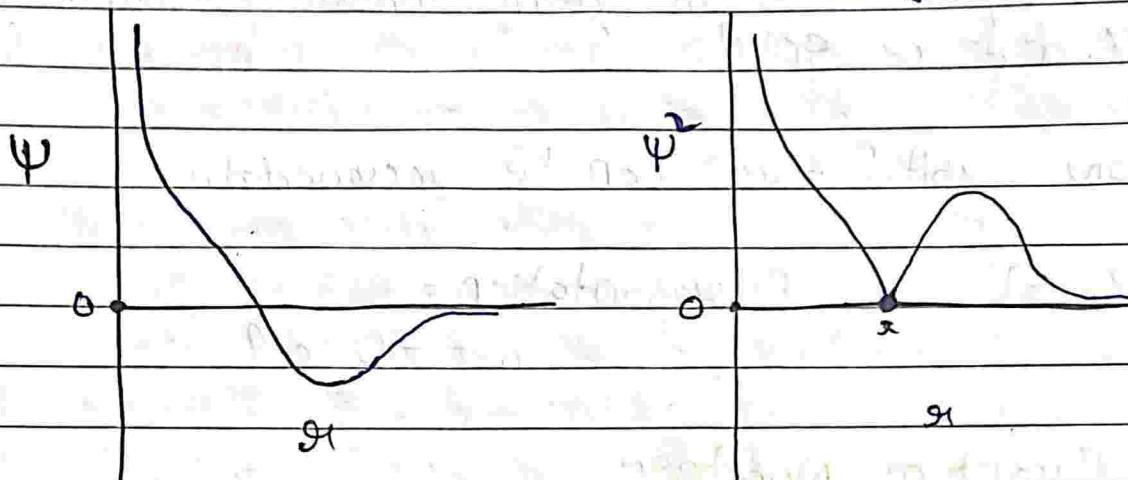
P-graph niche kar

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\* 2S orbital

x-axis ko cut according to node  
Krega.



⇒ Nodes :- "Area in an atom where probability of find  $e^-$  become zero ( $\psi^2 = 0$ )"

S-orbitals are spherical in shape: ○

① Radial Node :- Spherical area of node

$$\text{no. of Radial nodes} = (n - l - 1)$$

② Angular Node :- no. of angular ~~con~~ node = [l]  
(axPS)

# S-orbital is closest to nucleus, so  $e^-$  in S-orbital have highest effect at nucleus.

# P-orbital

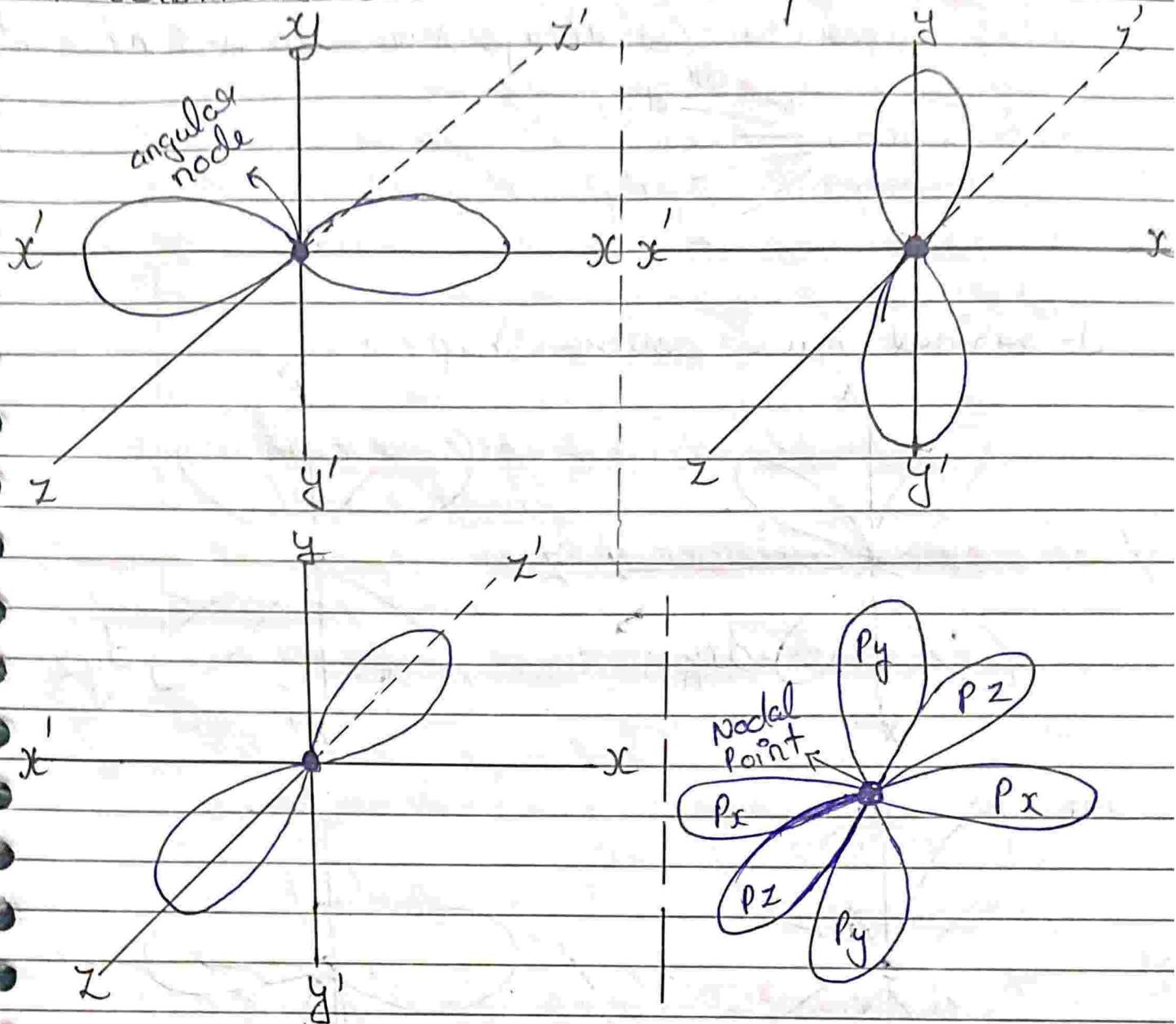
"from  $n=2$  and so on, we always found p-orbital."

$$l=1, m=3 \rightarrow -1(P_x)$$

$$l=1, m=0 \rightarrow 0(P_y)$$

$$l=1, m=1 \rightarrow +1(P_z)$$

P-orbitals are Dumbbell Shaped.



Total no. of nodes = Radial node + Angular nod

$$\begin{aligned} TN &= R \cdot N + A \cdot N \\ &= n - \ell - 1 + \ell \end{aligned}$$

$$TN = n - 1$$

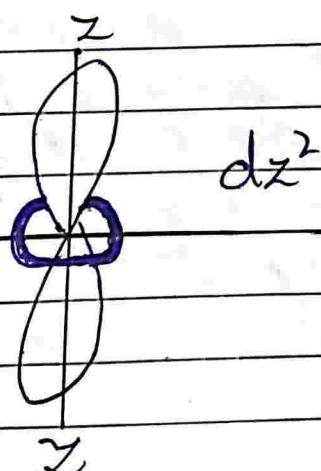
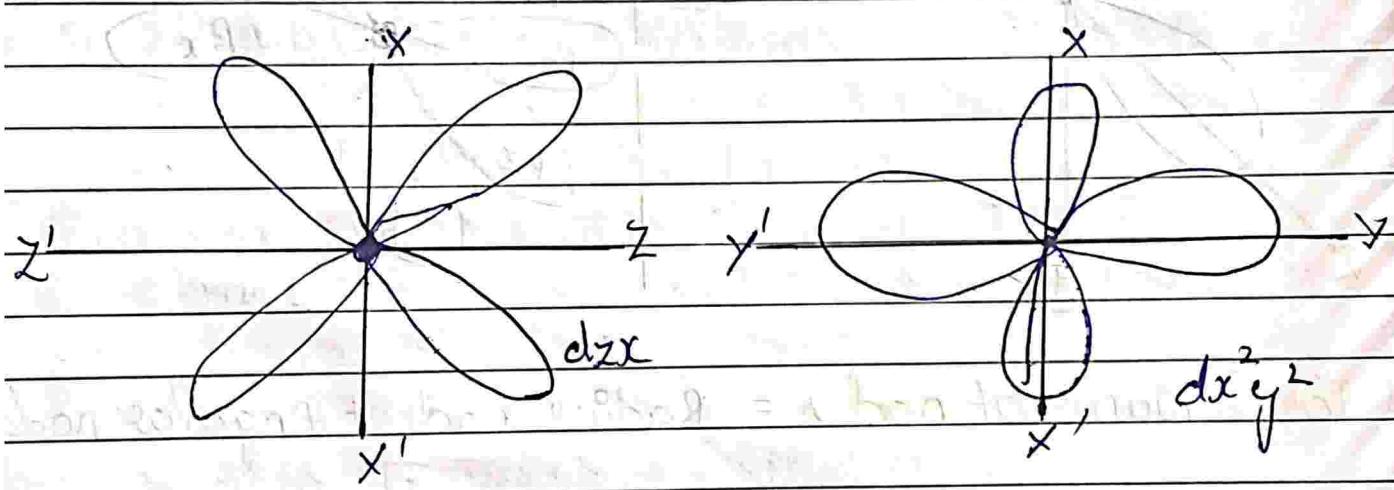
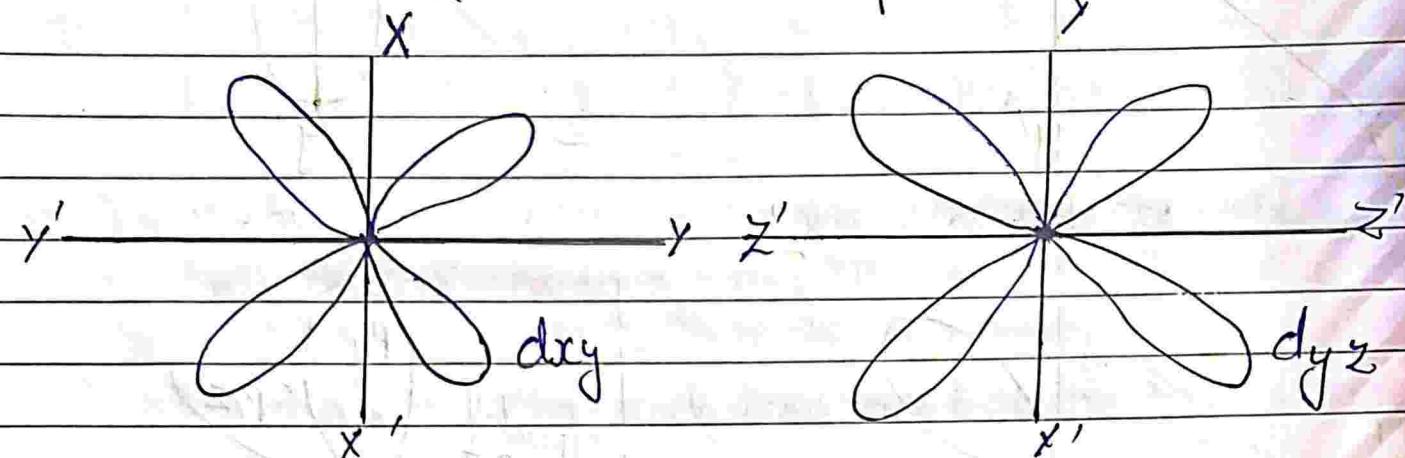
## Fd-orbitals.

"d-orbital is always seen from  $n=3$  and so on"

$$l=2, m=5$$

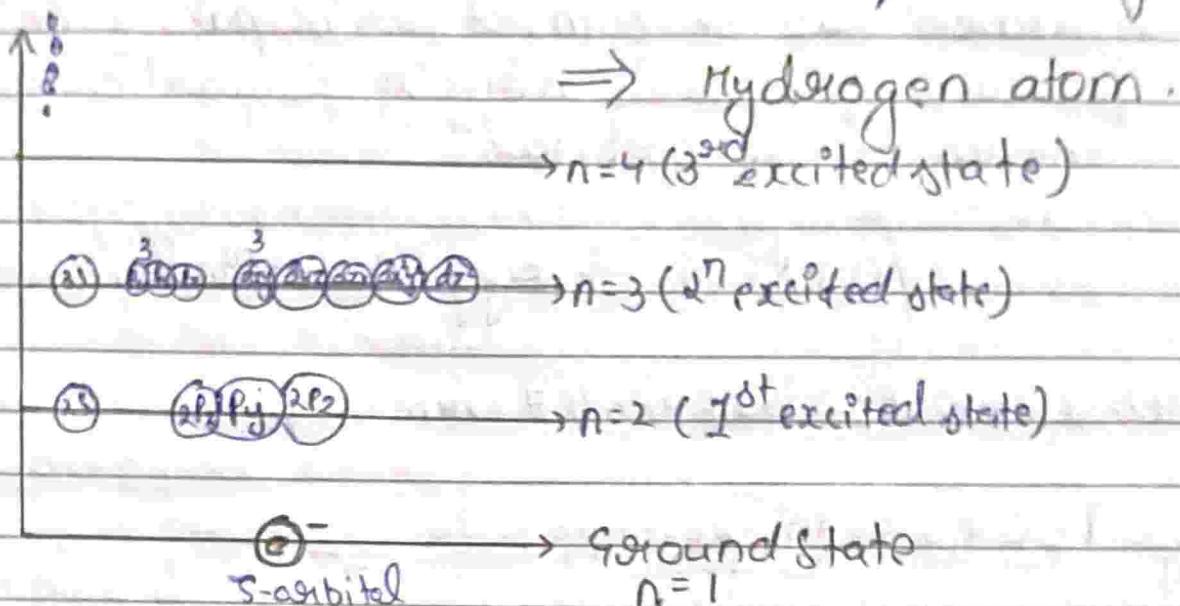
$d_{xy}$   
 $d_{yz}$   
 $d_{zx}$   
 $d_{z^2}$   
 $d_{x^2y^2}$

d-orbital have double shape



# # Energies of Orbitals ("Bohr and Bury concept")

- \* Discrete and degenerated orbitals :-  
"Orbital that have same or equal Energies"



"when we provided same energy to e<sup>-</sup> different times, sometimes it occupies s, some times p<sub>x</sub> or p<sub>y</sub> or p<sub>z</sub> or, say ..., of same shell"

"that means all the orbitals of same shell have same energy in hydrogen atom"

$$2s = 2p_x = 2p_y = 2p_z \Rightarrow \text{degeneracy}$$

degenerated orbitals

# Multi<sup>o</sup> e<sup>-</sup> Atoms.

"For multi-e<sup>-</sup> atoms we have to learn Bohr and Buoy Rule":-  
also known as (n+l) rule.

"It states that the sum of principal & azimuthal quantum number is directly proportional to Energy of the orbital".

$$[n+l \propto \text{Energy of orbital}]$$

Note:- f orbitals are total 7 are all are degenerated.

Energy:-  $1s < 2s < 2p_x = 2p_y = 2p_z < 3s \dots$   
degenerated.

Likewise:-  $dxy = dyz = dxz = dx^2y^2 = dz^2$   
are degenerated.

# Aufbau Principle:- "e<sup>-</sup> in an atom should be placed in increasing order of energy of the orbitals".

$\Rightarrow 1s \quad 2s \quad 2p$

$3s \quad 3p \quad 4s$

$3d \quad 4p \quad 5s$

$4d \quad 5p \quad 6s$

$4f \quad 5d \quad 6p \quad 7s$

$5f \quad 6d \quad 7p \quad 8s$

# Hund's Rule :- When  $e^-$  are placed in an orbital that the orbital should be first half filled & and followed by fully filling it, then move to higher energy orbital.

# Pauli	$\begin{array}{ c c } \hline 1L & 1L \\ \hline \end{array}$	$\begin{array}{ c c c } \hline 1L & 1 & 1 \\ \hline \end{array}$
	X	✓

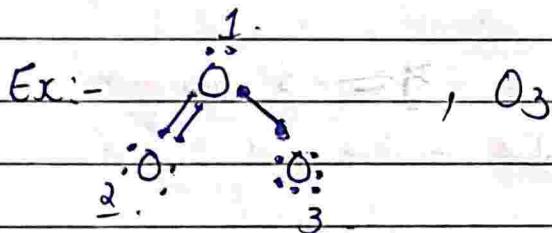
# Pauli's Exclusion Principle :- no two electron can have all 4 same quantum numbers. Because 2  $e^-$  in same shell, same subshell and same orbital but they will always have different spin.

## # Formal Charge on an Atom

$$F.C = [V - \frac{N.B - B}{2}] , V = \text{Valence } e^-$$

$N.B = \text{non-Bonding } e^-$

$B = \text{Bonding } e^-$



$$O\ 1^{\text{st}} \ F.C = 6 - 2 - \frac{1}{2}(6) = +1$$

$$O\ 2^{\text{nd}} \ F.C = 6 - 4 - \frac{1}{2}(4) = 0$$

$$O\ 3^{\text{rd}} \ F.C = 6 - 6 - \frac{1}{2}(x) = -1$$

# CHAPTER:- CHEMICAL BONDING ...

→ Every atom loses its energy by forming chemical bonds to achieve stability.

Energy of 1  
stability.

# Molecules :- These are gp. of atoms that are held together due to attraction and physical forces.

Types of Molecules :-

① Homonuclear Diatomic :- Two atoms of same nucleus (same atom) bond together.  
Ex:-  $H_2$ ,  $Cl_2$ ,  $F_2$ ,  $Br_2$ ,  $O_2$ ,  $I_2$

② Heteronuclear Diatomic :- Two atoms of different nucleus (different atom) bond together.  
Ex:-  $HCl$ ,  $HBr$ ,  $CO$ ,  $NO$

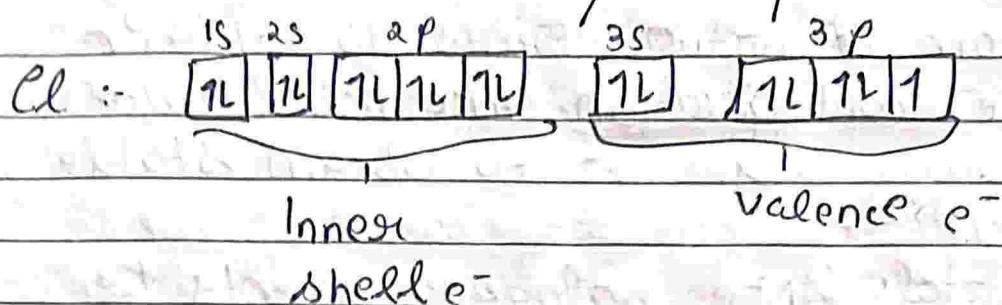
③ Homonuclear Polyatomic :- Many atoms of same nucleus bond together.  
Ex:-  $O_3$ ,  $S_6$ ,  $S_8$ ,  $P_4$

④ Heteronuclear polyatomic :- Many atoms of different type bond together.  
Ex:-  $NH_3$      $CH_4$ ,  $Cl_2$

## # Lewis Concept :-

:- Kernel: nucleus + Inner shell e<sup>-</sup>

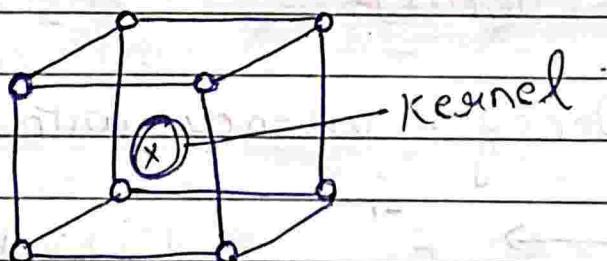
# Inner shell never participate in bonding.



⇒ "Every atom in nature tries to complete its valence shell i.e. to achieve 8 e<sup>-</sup> in the valence shell."

→ Either atoms donates/accepts/share-

⇒ We have to imagine atom as cube. Where the centre of the cube is called kernel and the 8 corners are occupied by e<sup>-</sup> of valence shell e<sup>-</sup>.



## \* Lewis Dot structure

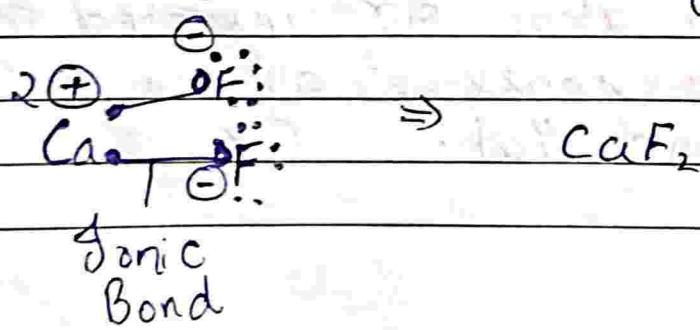
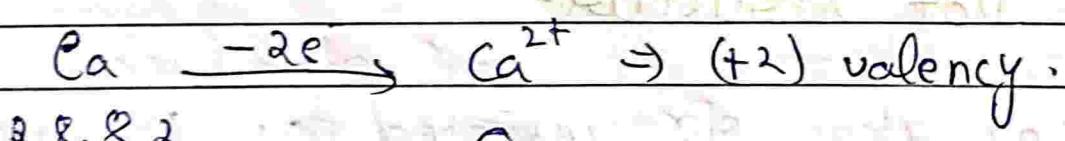
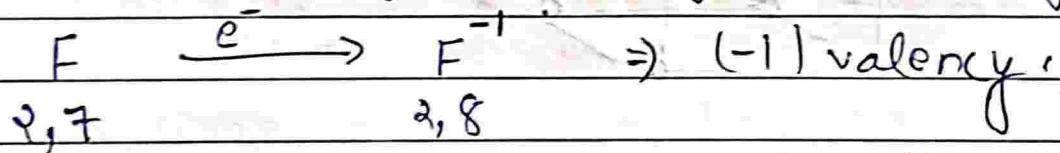
- Symbol of atom represented as kernel
- The valence shell e<sup>-</sup> are represented by dots around that.

Ex:- H, He, C, O, P

## # Kossel Concept :-

- Every atom in nature tries to obtain stable e<sup>-</sup> configuration like noble gases.
- Metals always loses e<sup>-</sup> to obtain stable E.C.
- Non-metals like halogens accept the e<sup>-</sup> to obtain stable E.C.
- \* When an atom loses its e<sup>-</sup>, it obtains a +ve charge & the atom accepts the e<sup>-</sup> it obtains negative charge.
- The +ve charged and -ve charged ions are attracted by electrostatic force of attraction & held each other and such bond are called Electrovalent Bond. (Ionic bond)

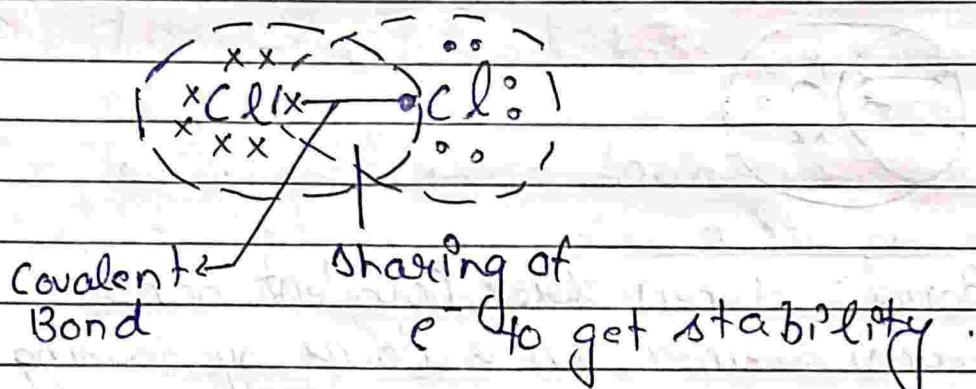
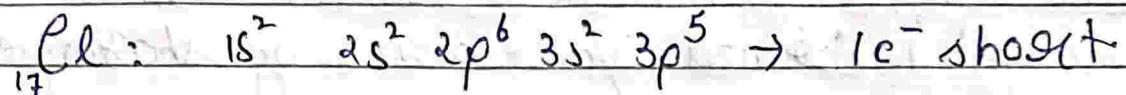
**Electrovalency :- Valency with charge-**



## # Octet Law:-

Every atom in nature tries to complete 8 e<sup>-</sup> in its valence shell either by accepting or donating the e<sup>-</sup> or by sharing of e<sup>-</sup>.

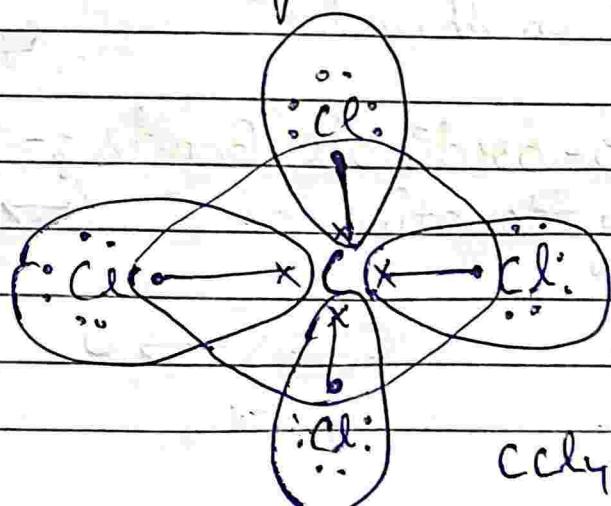
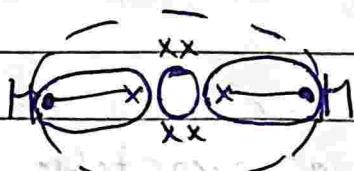
## # Langmuir :- (Co-valent Bonds)



Theory :- "When e<sup>-</sup> are mutually shared b/w two atoms in order to complete their valence shell the force of attraction/bonding is called Co-valent Bond."

\* e<sup>-</sup> must be shared mutually i.e. one e<sup>-</sup> each atom must involve in sharing.

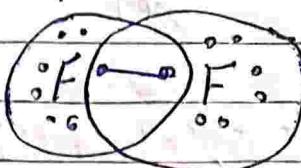
Ex:-



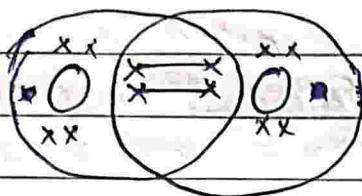
## Types of covalent Bond :-

① Single Bond :- When only one covalent bond is formed b/n 2 atoms by sharing of  $e^-$

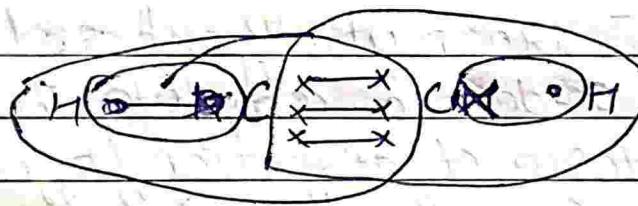
Ex:-



② Double Bond :- When two covalent bond (four  $e^-$ , two  $e^-$  each) formed b/n 2 atoms by sharing  $e^-$ .



③ Triple Bond :- When three covalent bond (6  $e^-$ , 3  $e^-$  each) formed b/n 2 atoms by sharing  $e^-$ .



$T$  = total valence  $e^-$

$V$  = valence  $e^-$

$S$  = shared  $e^-$

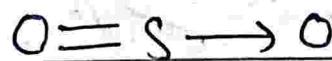
$U$  = unshared  $e^-$

$$S = T - V$$

$$U = V - S$$

# Co-ordinate Bonds :- The bonds that are formed by donating of  $e^-$  [→]

-  $SO_2$



$S = e^-$  donor

$O = e^-$  acceptor

## # Formal Charges

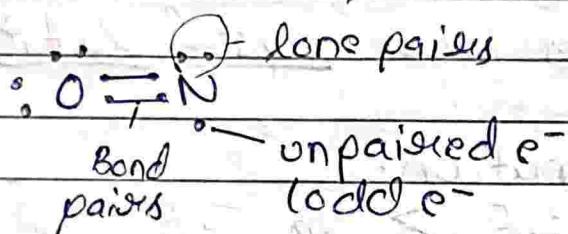
"used to determine the most stable structure of a molecule, the structure with least formal charge will be considered as most stable."

Formal charge = Total  $e^-$  in bonding - Non bonding  $e^-$  -  $\frac{1}{2}$   $e^-$  in valence shell

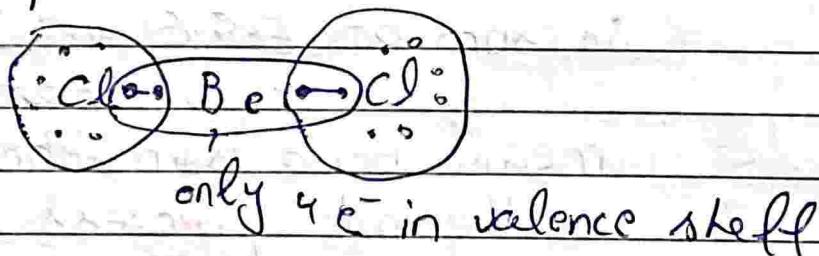
$$FC = T - NB - \frac{1}{2} S$$

## Limitation of Octet Law :-

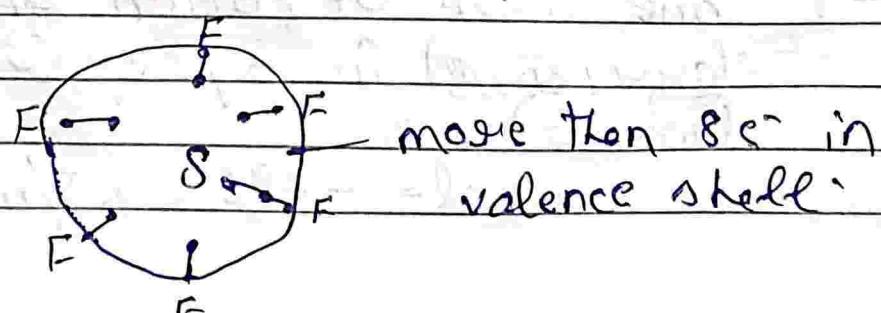
### ① odd $e^-$ molecules :-



### ② Incomplete valence shell :-



### ③ Expand Octet molecule :-



④ Nobel gases also form bonds and they do react.

Ex:-  $XeF_4$ ,  $XeF_6$ ,  $K_3XeF_6$

⑤ It does not deal with actual geometry of molecule and does not deals with PES ~~of~~ relative stability molecules.

# Ionic Bonds or Electrovalent Bond.

"Bonds formed due to the complete transfer of  $e^-$ "

- Ease of ionic bond formation:-
- Ease with which an atom loses its  $e^-$  to form a cation & the ease with which an atom gains an  $e^-$  to form anion.
- Ease with which these ions combine and arrange themselves in a crystal lattice.

In bond formation:

$$E.G.E > \text{I.E}$$

(with  
(-))

So, sum of E.G.E and I.E should come out as negative means bond formations are exothermic process.

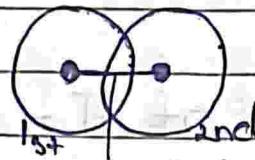
# Lattice Enthalpy :- "Energy ~~(released)~~<sup>(required)</sup> when  $\pm$  one mole of cation and anion ~~(arranged)~~<sup>(break)</sup> in a proper lattice."

$$NaCl = -786 \text{ kJ/mol}$$

# # Bond Parameters

① Bond length :- It is the equilibrium distance between the nucleus of two atoms.

Covalent Bond:-

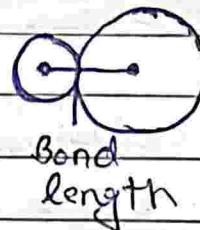


Bond length.

$$\text{Bond length} = r_1 + r_2$$

# X-Ray Diffraction method is used find Bond length.  
(and Spectroscopy method)

Ionic Bond:-



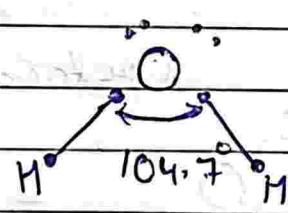
$$\text{Bond length} = r_c^+ + r_a^-$$

Bond length  $\propto$  1 / Bond order  $\propto$  1 / Bond dissociation Energy

② Bond Angle :- Angle between two bonds pairs of an central atom.

$B \cdot A \propto$  Electro-ve

$B \cdot A \propto \frac{1}{\text{size}}$

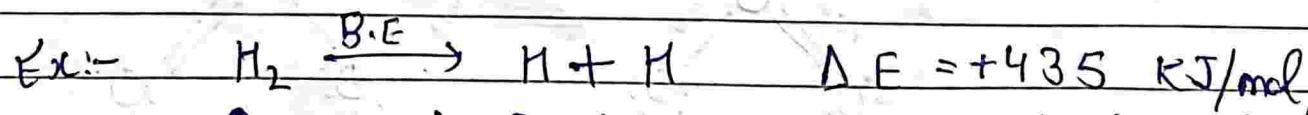


$$Sp_3 \rightarrow 109^\circ$$

$$Sp_2 \rightarrow 120^\circ$$

$$Sp \rightarrow 180^\circ$$

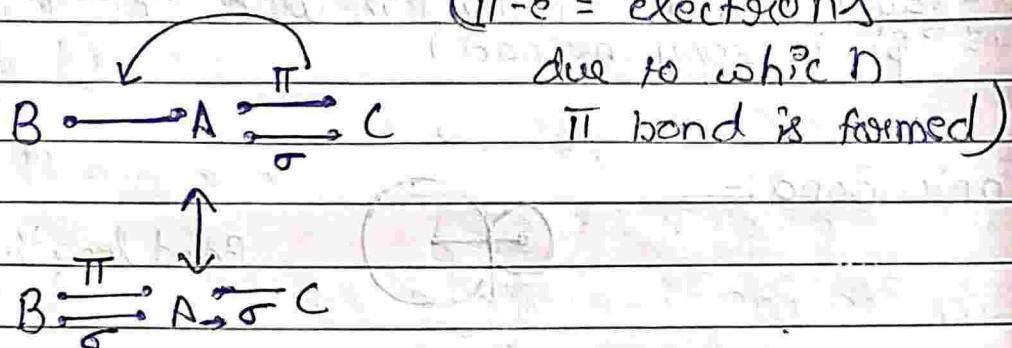
③ Bond dissociation Energy :- It is the energy required to break one mole of bonds.



④ Bond Order :- No. of bonds present b/w two covalently bonded atoms  
It may be :-  $1, \frac{1}{2}, 3, \frac{1}{2}, \frac{3}{2}, \frac{5}{2}$  etc.

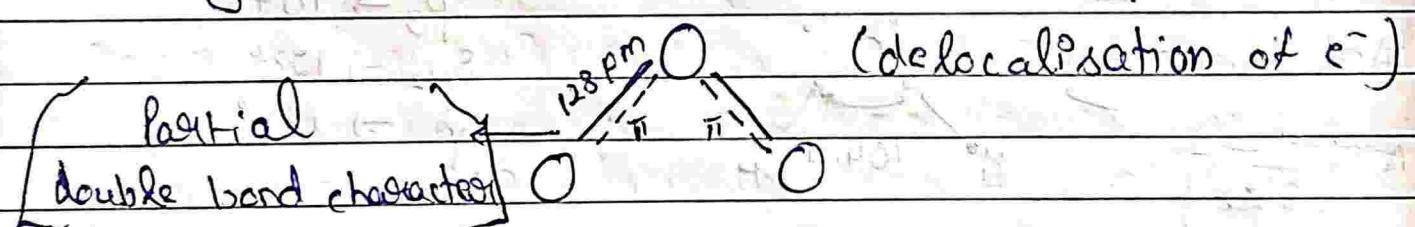
## # Resonance

It is the delocalization of  $\pi$ -bonds ( $\pi-e^-$ )  
It is shown by ( $\longleftrightarrow$ ) double headed arrow.

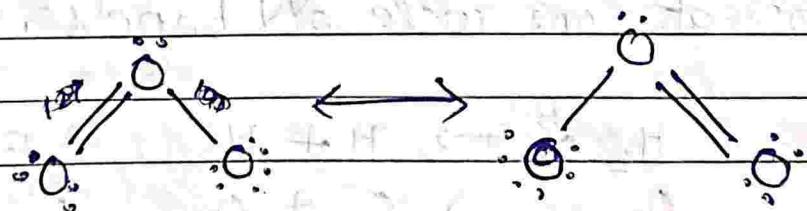


Canonical Structures :- structures which have same energy, position of atoms, non-bonding & bonding  $e^-$  are taken as canonical form. (They are imaginary)

Resonating actual structure of  $O_3$  (Hybrid)



Canonical structures :-



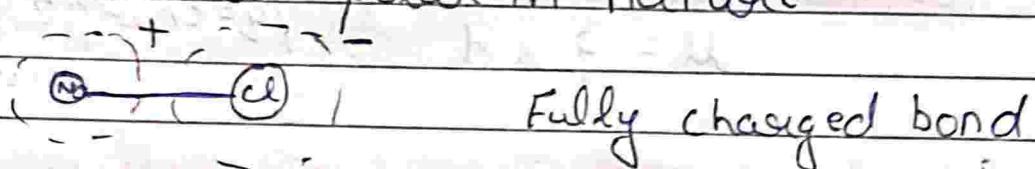
\* Canonical forms have higher energy than actual hybrid structure

# Hence Resonance stabilizes the molecule by decreasing its energy.

## # Polarity of Bonds

"The charge separation on the bonded atom that is on one side (partially or fully) +ve and another -ve charge ~~is~~ are present then such bonds are called polar bonds"

\* All ionic bonds are polar in nature.



⇒ Polarity of covalent bonds :-

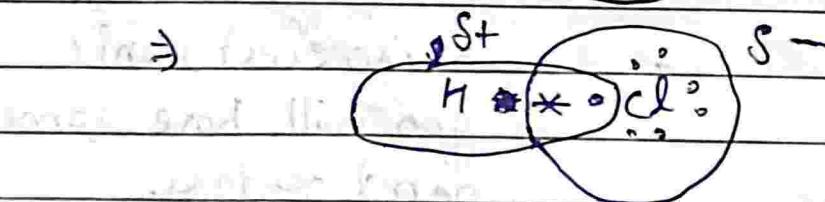
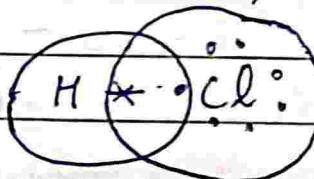
- Homonuclear molecule (Non-polar)
- Heteronuclear molecule (Polar)

Heteronuclear molecules have different atoms unlike homonuclear molecules, so the atoms have different electronegativity hence the possess polar nature.  
[Note:- Not all heteronuclear molecules are polar]

Conditions:-

① Difference Observable difference in Electronegativity  
 $\text{CH}_4 \rightarrow$  non polar C & H almost same E.N

Ex:-  $\text{HCl}$ .



⇒ when a more E.N atom is covalently bonded with an atom with less E.N, then the atom with more EN will partially attract the shared paired e<sup>-</sup> towards itself which gains s- and atom with less EN gain s+ charge, such covalent bonds are called Polar covalent bonds.

# Polarity can be measured in terms of dipole moment (M)

$$M = q \times d \quad \left[ \begin{array}{l} q = \text{charge magnitude} \\ d = \text{distance b/w atom} \end{array} \right]$$

S.I unit :- cm

Common unit :- Debye =  $3.33564 \times 10^{-30}$  cm

$$nD = n \times 3.33564 \times 10^{-30} \text{ cm}$$

# In dipole moment between q & d, d is dominating character

Bond order trick.

$$1 \rightarrow 12 - 2$$

$$\text{ex:- oxygen } - 8e^- \quad 1 \rightarrow 14 - 3$$

$$\text{No. of total - Bond} \quad \rightarrow 16 - 2$$

$$\cdot \text{electrons} \quad \text{order.} \quad \rightarrow 18 - 1$$

$$\rightarrow 20 - 0$$

$$\rightarrow 2 - 1$$

$$\rightarrow 4 - 0$$

$$\rightarrow 6 - 1$$

$$\rightarrow 8 - 0$$

$$\rightarrow 10 - 1$$

# iso electronic

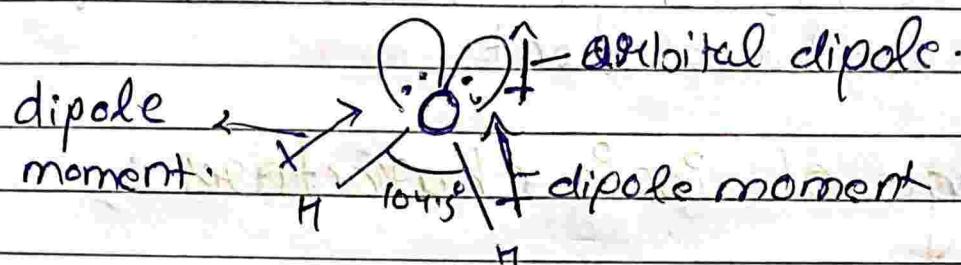
gps will have same bond order.

\* Dipole moment is a vector quantity.

$$\begin{bmatrix} +\rightarrow \\ +ve \quad -ve \end{bmatrix}$$

\* For non-polar compound  $\mu = 0$

\* orbital dipole (orbital polarity) :- dipole moment because of the lone pair of e<sup>-</sup>



### # FAJAN's Rule:-

"No covalent bond is 100% covalent because it consists some ionic characters and no ionic bond is 100% ionic because it consists some covalent character."

⇒ Even Homonuclear like H<sub>2</sub>, which looks non-polar but actually have slight to very slight ionic character.

⇒ The power of attracting the e<sup>-</sup> cloud of an ion towards itself is c/s Polarising Power.

⇒ The extent to which an anion gets polarised is c/s Polarisability.

Note:- Smaller the size of cation, more will be polarising power, more will be covalent character. Or the larger the amt of charge on cation.

Note:- Larger the size of anion, larger will be the polarisability, Larger will be covalent character.

⇒ If two cations of equal charge & size then the cation with pseudo noble gas config. will have high polarising power.

Pseudo noble gas config.: -  $(n-1)d$  orbital. simply a-d-orbital element.

# Percentage of ionic characters :-

$$\Rightarrow \frac{\text{Given dipole moment}}{\text{net dipole moment}} \times 100$$

# VSEPR Theory

(Valence shell Electron Pair Repulsion Theory)

# This theory deals with the shape of the molecule.

# By :- Sidwick and Powell & updated by :- Nyhan & Gellispi.e.

⇒ Postulates :-

- The shape of molecule will depend on bonding and non-bonding  $e^-$  pair present in the valence shell of central atom.

- As the  $e^-$  pairs are all negatively charged hence will repel each other in the valence shell.
- So, in order to reduce the repulsion these  $e^-$  pairs will tend to move away from each other in such a way that they will have minimum repulsion b/w them.
- As we consider the shape of central atom & spherical the  $e^-$  pairs arrange themselves around the spherical central atom. So that they can achieve maximum stability / max distance.
- As per VSEPR theory a double bond and triple bond are considered as a super single bond.
- It is applicable to all the canonical or resonating forms.

Updated Postulate :-

- The lone pair of  $e^-$  belongs only to the central atom but bond pair of  $e^-$  belongs to both atoms. So, lone pair have more repulsion as compared to the  $B-P-e^-$ .

$\Rightarrow$  Repulsion Order:-

$$l \cdot p - l \cdot p > l \cdot p - B \cdot p > B \cdot P - B \cdot p$$

↓    ↓

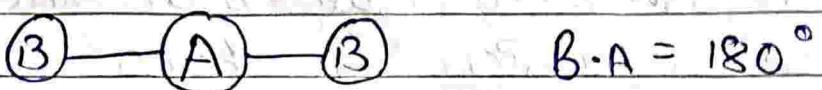
maximum    minimum

repulsion    repulsion.

# A/C to VSEPR Theory types of molecules.

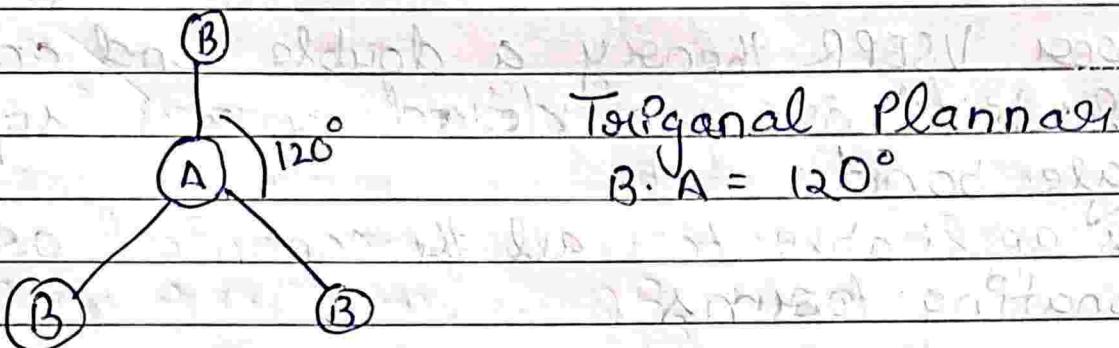
## ① Molecules without Lone Pairs :-

(i) 2 Bond pairs ,  $AB_2$  type , Ex:-  $\text{BeCl}_2$

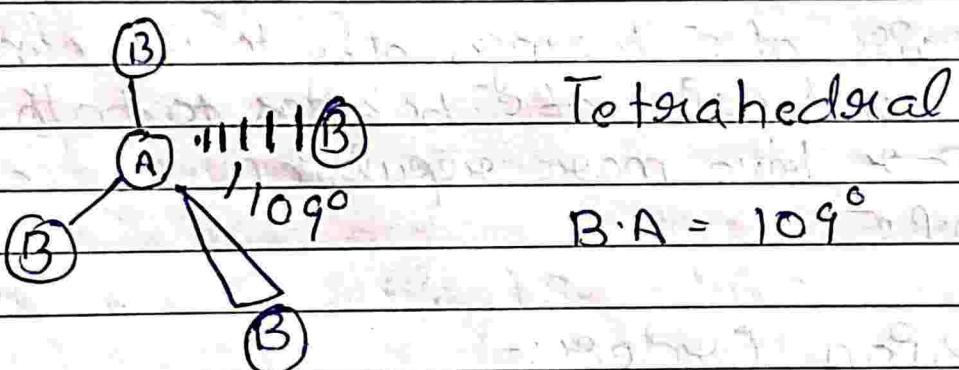


Linear

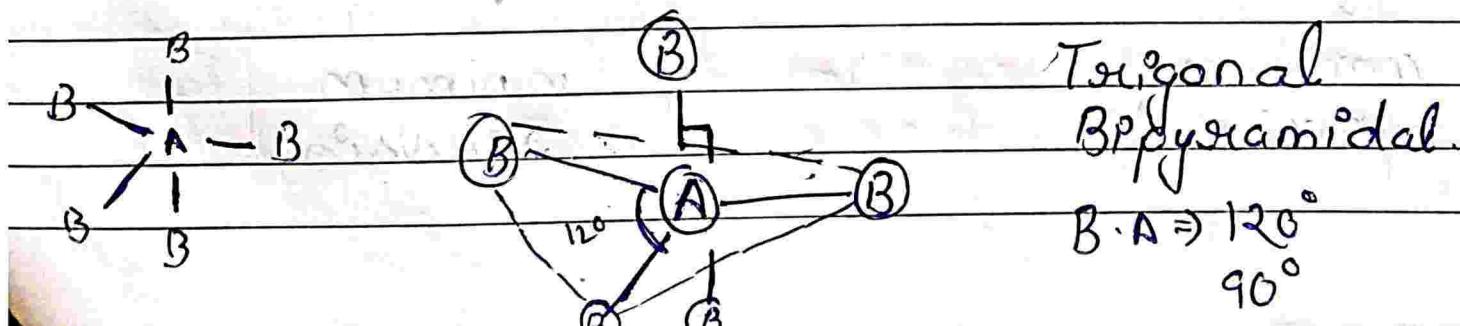
(ii) 3 Bond pairs ,  $AB_3$  type , Ex:-  $\text{BCl}_3$



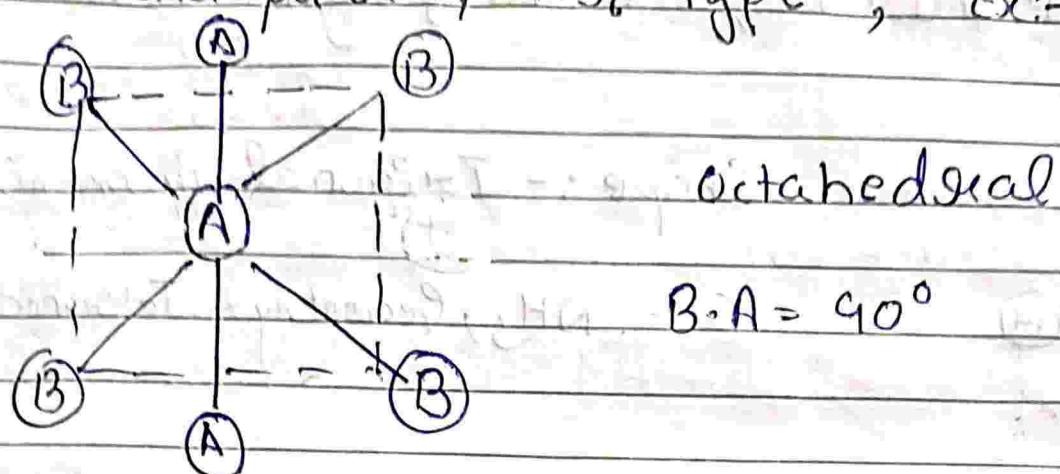
(iii) 4 Bond pairs ,  $AB_4$  type , Ex:-  $\text{CH}_4$



(iv) 5 Bond Pairs ,  $AB_5$  , Ex:-  $\text{PCl}_5$

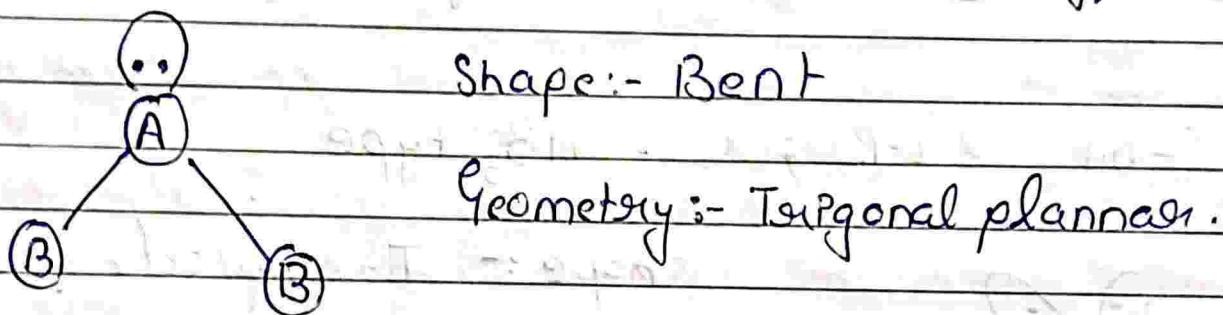


(v) 6-Bond pairs,  $AB_6$  type, Ex:-  $SF_6$

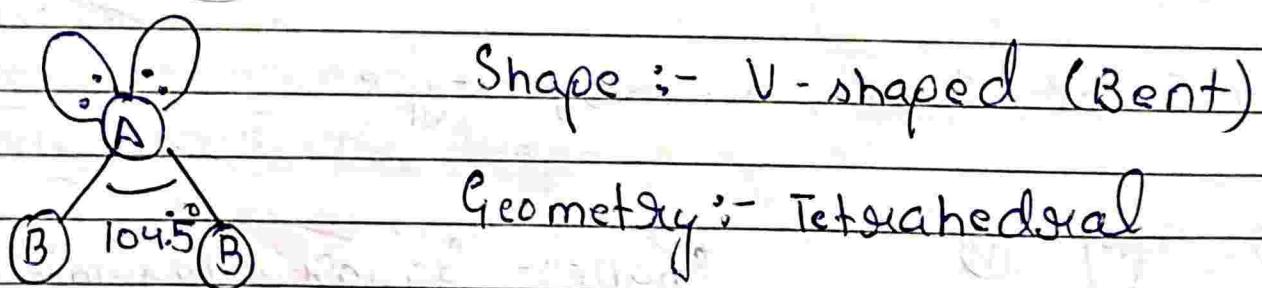


② Molecules with Lone-Pair and Bond-Pairs  
Let  $x$  = lone pair.

(i) 2-Bond Pairs; 1-lone Pair;  $x, AB_2$  type

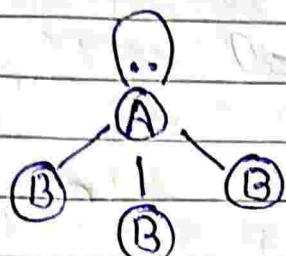


(ii) 2-B.P.; 2 L.P. ;  $x_2, AB_2$  type



# Shape is determined by lone pairs.

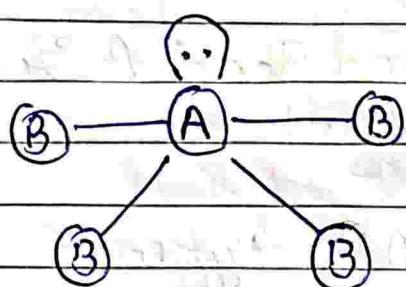
(iii) 3-B.P, 1 L.P ; X,  $AB_3$  type



Shape :- Trigonal Pyramidal

Ex:-  $\text{NH}_3$ , Geometry:- Tetrahedral

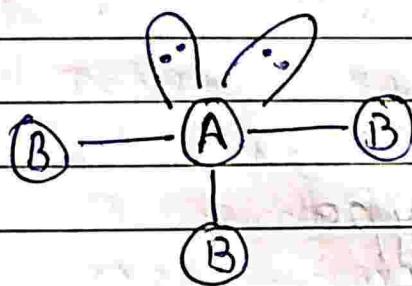
(iv) 4-B.P, 1 L.P ; X,  $AB_4$  type



Shape :- See-Saw

Geometry :- Trigonal bipyramidal

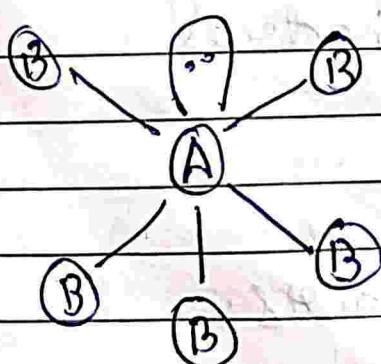
(v) 3-B.P, 2 L.P ; X,  $AB_3$  type



Shape :- T-shaped

Geometry :- Trigonal Bipyramidal

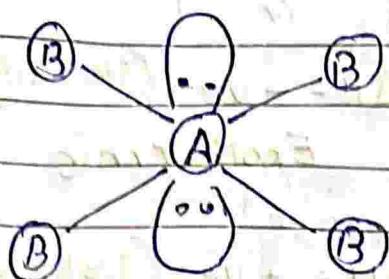
(vi) 5-B.P, 1 L.P, X ;  $AB_5$  - type



Shape :- Square Pyramidal

Geometry :- Octahedral

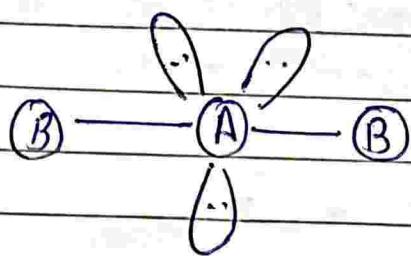
(vii) 4-B.P, 2-L.P ;  $x_2$ ,  $AB_3$  type



Shape :- Square Planar.

Geometry :- Octahedral.

(viii) 3-B.P, 3-L.P,  $x_3$ ,  $AB_2$  type.



Shape :- Linear

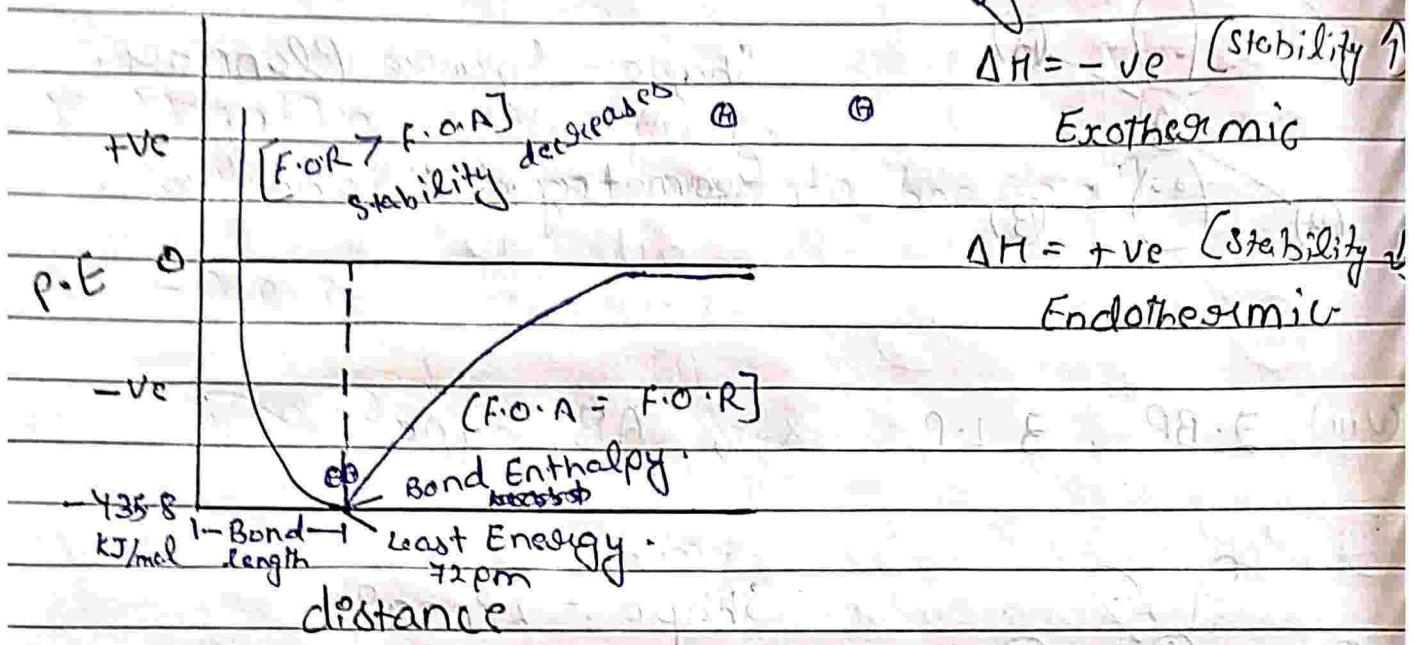
Geometry :- Trigonal  
Bipyramidal

## # VBT (Valence Bond theory).

"it is based on wave mechanics and quantum theory"

- Initially both the H-atom are away from each other, no force is acting between them.
- Now as both the atoms starts approaching each other the force of attraction starts increasing.
- As the atoms gets closer to each other the new force i.e. force of Repulsion will starts acting.
- finally both the atoms stabilizes each other in such a way that  $F.o.A = F.o.R$  and molecule formation takes place.

# Distance Vs Potential Energy Graph



## Important Points of Graph :-

- As the H atoms are away from each other potential energy will be zero.
- As  $F.O.A$  increases the atoms comes closer to each other and distance decreases  $\Rightarrow$  potential energy decreases.
- When  $F.O.A = F.O.R$ ; the  $a_1$ , point on the graph is the bond length & potential will be noted on y-axis, potential Energy and it is called Bond enthalpy, max energy released.
- As further tries to decrease the bond length the  $F.O.R$  increases which leads in a sudden increase of ~~pot~~ in potential Energy.
- For  $H_2$ ,  $a_1 = 72 \text{ pm}$  and Bond enthalpy  $P.E = -435.8 \text{ kJ/mol}$ .

## # Overlapping of Orbitals

→ There are 3 types:-



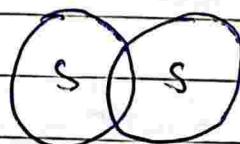
overlapping

- Positive overlapping.
- Negative overlapping.
- Zero overlapping.

### \* Positive overlapping :-

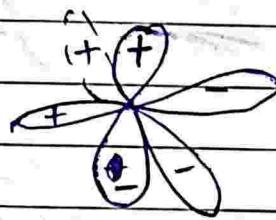
"It promotes molecule formation"

#### (i) S-S overlapping.



# S-orbital is spherical and can do overlapping from any direction.

#### (ii) # Directional Nature of orbital.



\* This +, - do not have any physical significance it only used mathematically to identify the direction as.

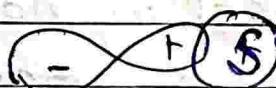
Phases of  $P_x, P_y, P_z$ .

#### (iii) + lobes of P orbital overlapping.



Sideways

#### (iv) S-P overlapping.



This is sideways overlapping

#### (v) P-P sideways overlapping

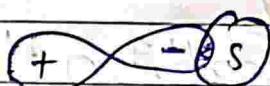


#  $\vec{e}_1$  in the overlapping must have opposite spine.

\* Negative overlapping.

"It do not promote bonding".

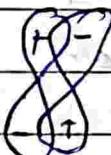
①  $-p \& s$  overlapping.



②  $+p \& -p$  overlapping.



③  $+p \& -p$  head on overlapping.



\* Zero overlapping:-

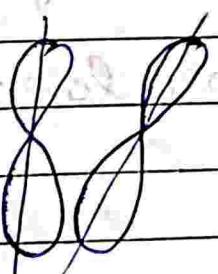
When there is no overlapping.

Ex:- ①



P-S sides head on.

②  $p_x \& p_z$  or  $p_y$



#  $p_x, p_z, p_y$  being directional in nature  
for molecule formation will only overlaps as,

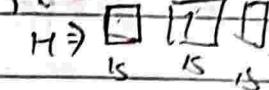
$p_x-p_x, p_y-p_y, p_z-p_z$

# Bonding structure of  $\text{CH}_4$  as overlapping By VBT

$\text{C}_{\text{ground}} = [1L] \quad [1L] \quad [1L] \quad [1L]$ ,  $\text{C}_{\text{excited}} = [1L] \quad [1L] \quad [1L] \quad [1L]$

state              1s            2s            2p            3s

$L = \text{hydrogen e}^-$



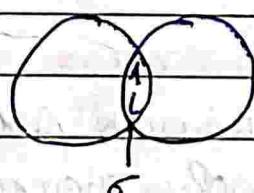
when we try to draw the angle b/w bonds come 90° but it should some  $109^\circ$  as per VSEPR Theory as overlapping concept does not work on poly atomic molecules it only works in diatomic Hence Pauli updated and Amended VBT to poly atomic molecules.

# Types of Overlapping & covalent Bond :-

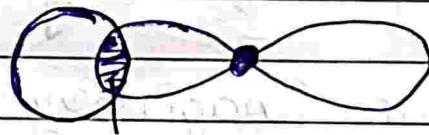
(i) Sigma Bond ( $\sigma$ ) :-

"Single covalent Bond"

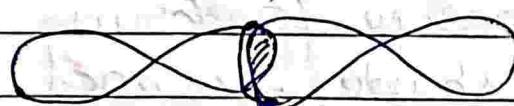
(a) S-S overlapping



(b) S-p head on overlapping



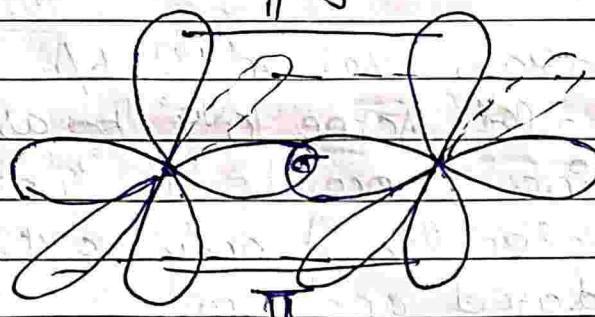
(c) P-P head on overlapping



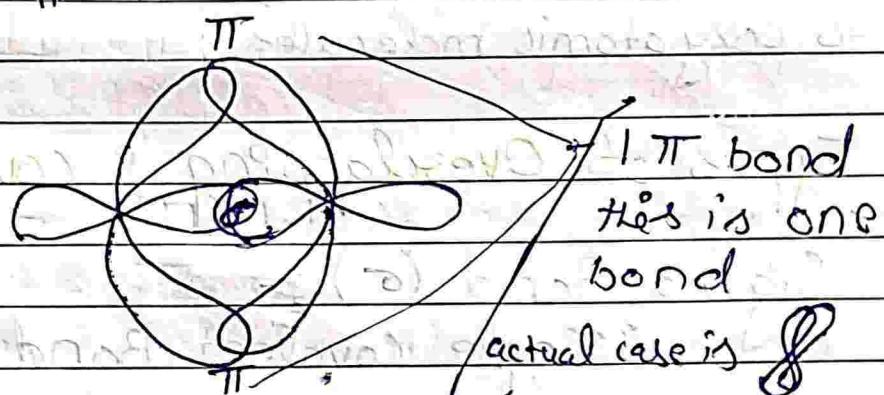
### (ii) Pie bond ( $\pi$ ) :-

"All the bonds formed after single sigma bonds (all double & triple bond)"

\* P-p over lapping



Actual:-



## # Hybridization

By :- Paulings

"It is an imaginary concept; It was proposed only to explain the bonds in poly atomic molec"

⇒ Hybridization is the intermixing of the orbital of the central atoms which have diff shapes and almost comparable energy to form new orbitals of equal shape, size and Energy.

# Isolated atoms hybridization do not takes places, it only takes places when an atom approaches for bonding.

## # Shape of hybrid orbitals :-



## # Conditions & Features

- Orbitals having almost same energy can only undergo hybridization.
- Orbitals of valence shell or penultimate shell can go under hybridization.
- Orbitals of only central atom goes under hybridization.
- The no. of ~~one~~ orbital going under hybridization are equal to new hybridized orbital formed.
- All the newly formed hybrid orbitals have same energy, size and shape.
- + • Hybridization do not require excitation ~~is~~ of  $e^-$  in atomic orbital
- Because of  $e^-$  repulsion all the new hybrid orbitals occupies the space in such a way that repulsion will be minimum.
- Empty orbitals can't go under hybridization.
- Hybridization is only used for  $\sigma$  bonds.  
Hence ~~force~~ as previously p-p overlapping is required for  $\pi$  bonds.
- Imp.** • Bond by hybrid orbitals is always stronger and stable than atomic orbital

## # Formula to find Hybridization.

$$H = n(\text{attach}) + L.P \text{ of that atom}$$

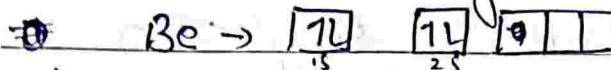
$n$  = no. of atoms attached to central atom.

## # Types of Hybridization

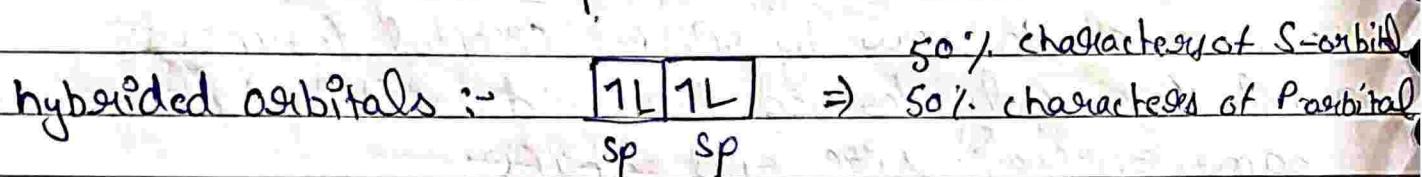
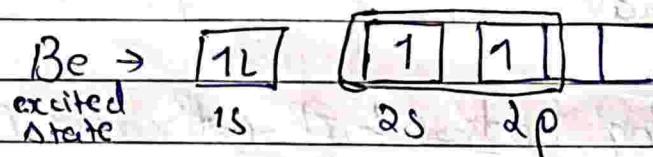
### ① ~~Simple~~ Sp hybridization.

Ex:-  $\text{BeCl}_2$

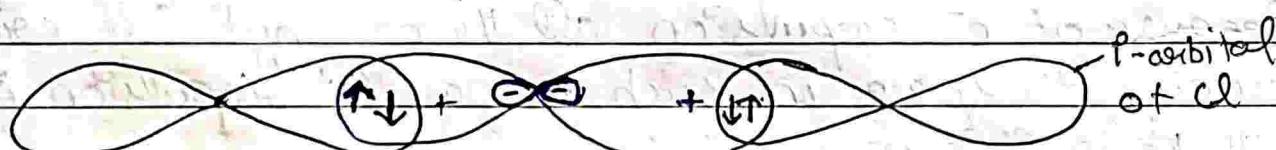
# Be is least e<sup>-</sup> negative so it will be a central atom.



→ when Cl<sup>-</sup> will approach Be one of the e<sup>-</sup> in Be get excited to 2p, for bonding there is always need of unpaired e<sup>-</sup>



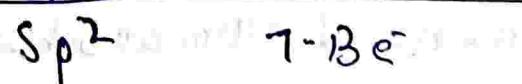
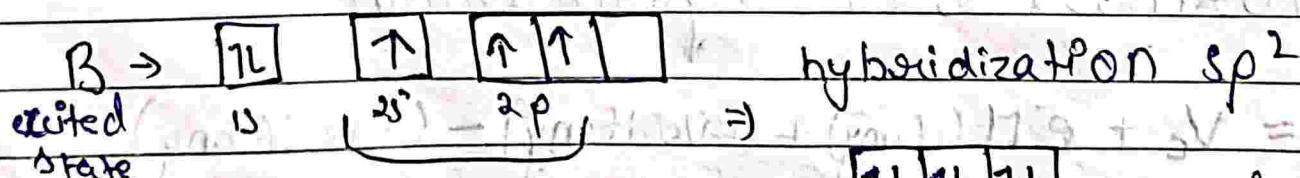
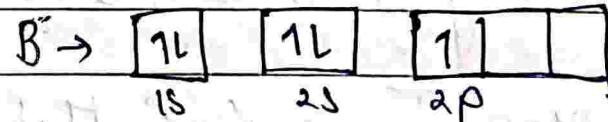
Shape :- Linear, Sp-p overlapping.



\* More the % character of S-orbital more will be the attractive nature of O-central atom.

### ② $\text{Sp}^2$ hybridization

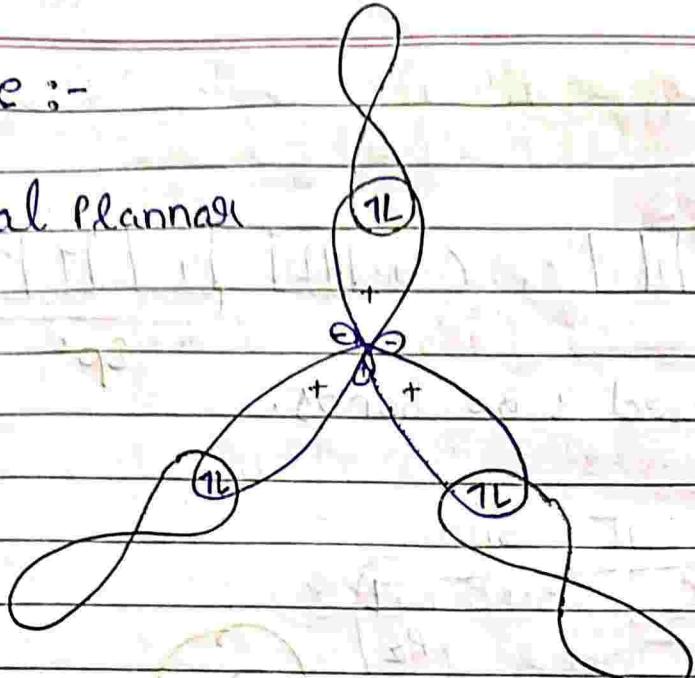
Ex:-  $\text{BCl}_3$



Shape :-

Trigonal Planar

$3P_z$  of Cl

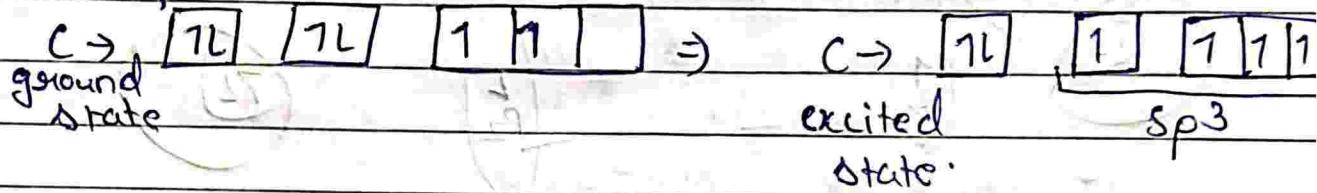


% S orbital  $\rightarrow$  33%

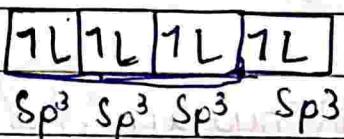
% P orbital  $\rightarrow$  66-7

③  $Sp^3$  hybridization.

Ex :-  $CH_4$



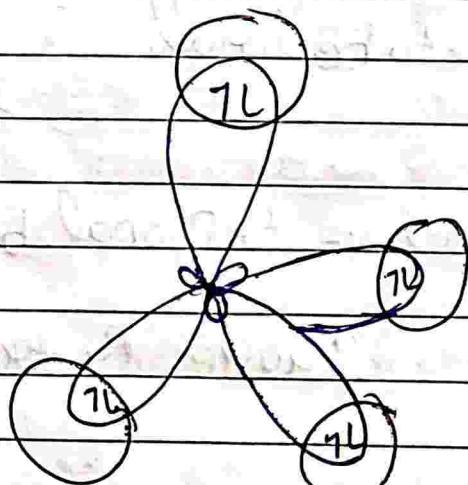
Hybridization  $Sp^3$



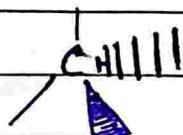
% S character = 25%  
% P character = 75%.

Shape :-

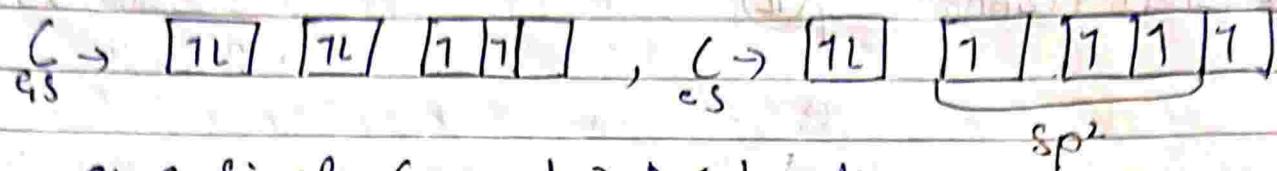
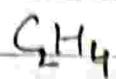
Tetrahedral.



actual :-

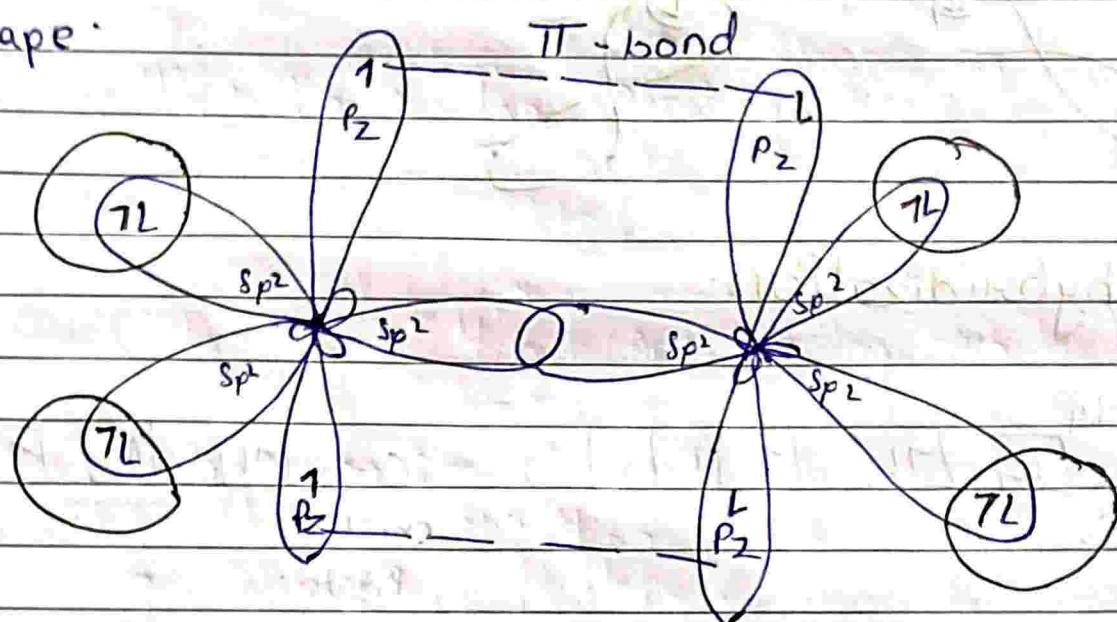


#  $\pi$  bond example:-



as a single C need 3  $\sigma$  bonds.

Shape:-



# Other:-

\*  $\text{sp}_3\text{d}$  ( $d_z^2$ )  $\rightarrow$  Trigonal Bipyramidal

\*  $\text{sp}_3\text{d}^2$  ( $d_{xy}, d_z^2$ )  $\rightarrow$  Octahedral

\*  $\text{dsp}^3$   $\rightarrow$  mostly:- trigonal bipyramidal

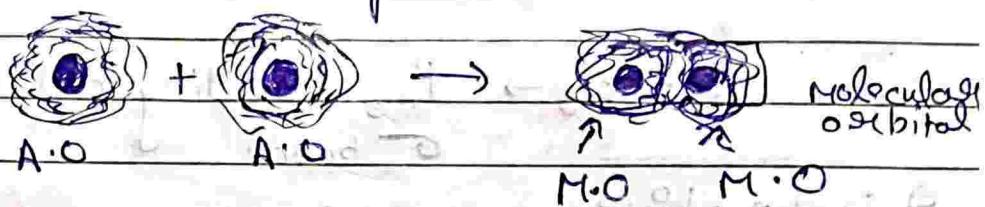
Square Pyramidal:-

\*  $\text{sp}_3^3\text{d}_3$   $\rightarrow$  Square Bipyramidal  
 $(d_{z^2}, d_{y^2}, d_{z^2}, d_{xy})$

# # Molecular Orbital Theory ...

By:- Hund's and Millikan

- \* The atomic orbitals of combining atom mix with each other completely for a new orbital.



- \* The  $e^-$  in atom was in A·O; but after mixing, now the  $e^-$  are found in M·O in a molecule.

- \* The atomic orbitals of nearly same energy can only form M·O by mixing.

- \* When  $e^-$  is in A·O it is influenced by only 1 nucleus hence it is called monocentric, but when  $e^-$  is in M·O it is influenced by more than one nucleus hence called polycentric.

- \* That the number of Atomic orbital mixed is equal to number of molecular orbital formed.

- \* The first orbital formed is called Bonding MO and other is called Antibonding MO.

- \* Bonding orbital :- less Energy, Max stability.  
Antibonding orbital :- More Energy, less stability.

Bonding Representation :- 1s, 2s, 3s ...

Antibonding MO Representation - 1s\*, 2s\*, 3s\*

## # Linear Combination of Atomic Orbital (LCAO)

$\Psi_A$  :- A.O of atom A

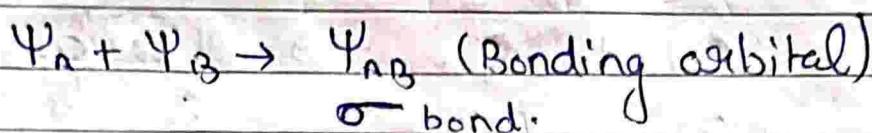
$\Psi_B$  :- A.O of atom B

linearly  
combine

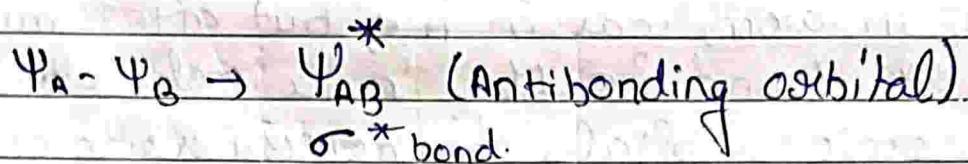
constructively

Destructively

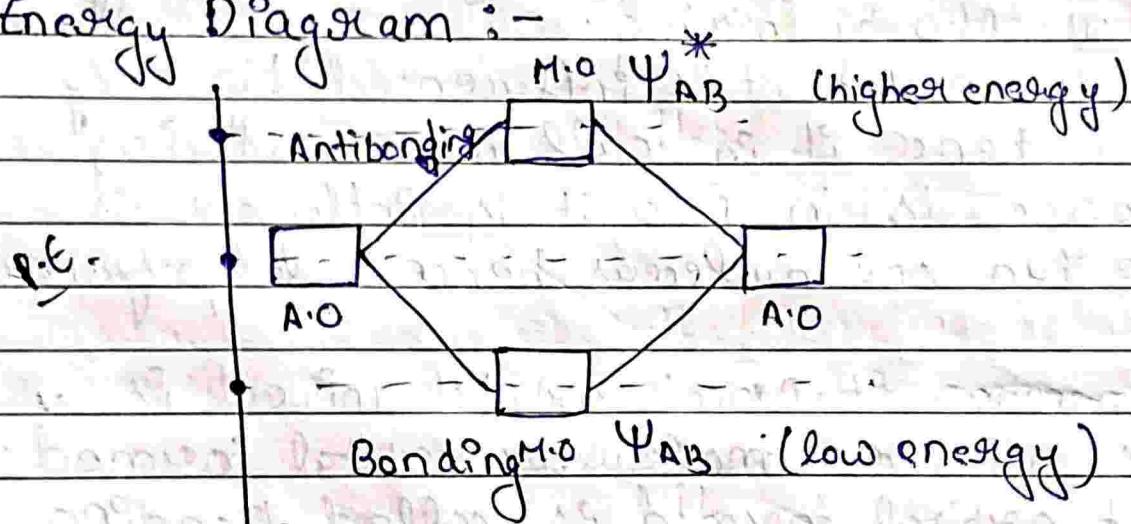
$\Rightarrow$  constructive



$\Rightarrow$  Destructive

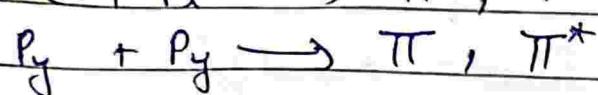
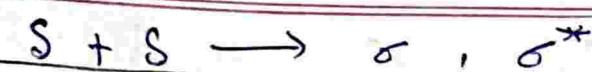


Energy Diagram :-

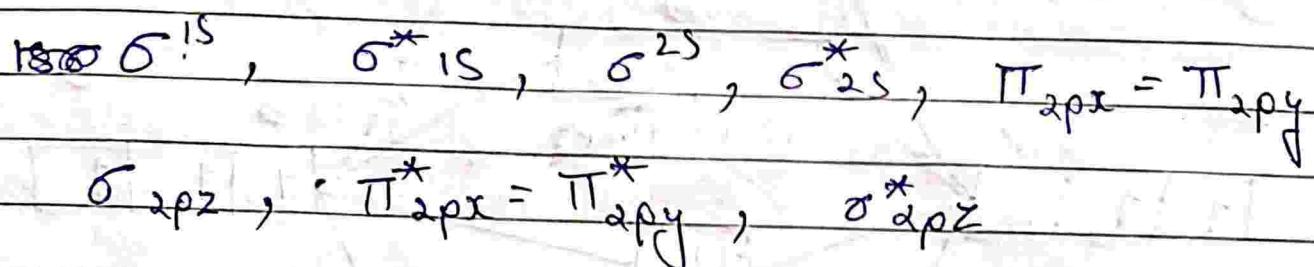
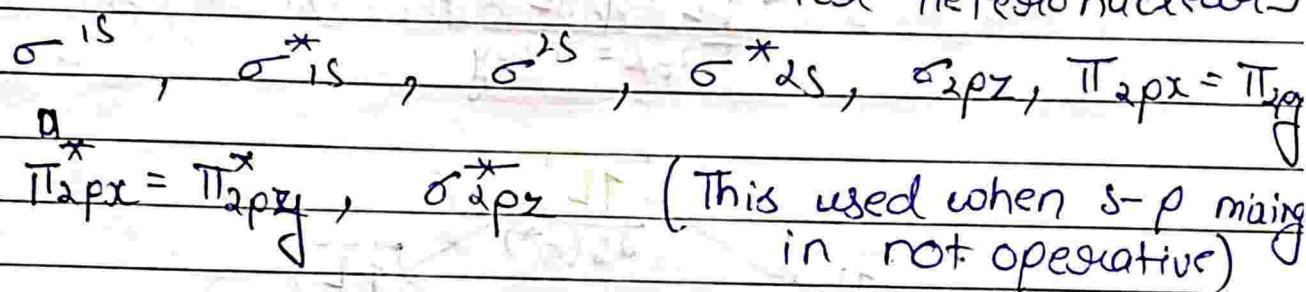


On combining of two A.O orbital formed M.O have most e<sup>-</sup> density ~~to off the center of the~~ of the is in b/w the nucleus to nullify repulsion b/w them.  
 This is constructive combine.

In case of destructive combine the e<sup>-</sup> density will be behind both nucleus which increases the repulsion.



(14 included)

# Upto 14 e<sup>-</sup> in a molecule order will be:-# Above 14 e<sup>-</sup> molecule : (But this not applicable for Heteronuclear) $N_A \Rightarrow$  No. of Antibonding e<sup>-</sup> $N_B \Rightarrow$  No. of Bonding e<sup>-</sup>
 $N_A = N_B \quad ]$  No molecule formation  
 $N_A > N_B \quad ]$  formation

$$\boxed{B.O = \frac{1}{2} [N_B - N_A]}$$

 $N_B > N_A \Rightarrow$  Molecule formation.

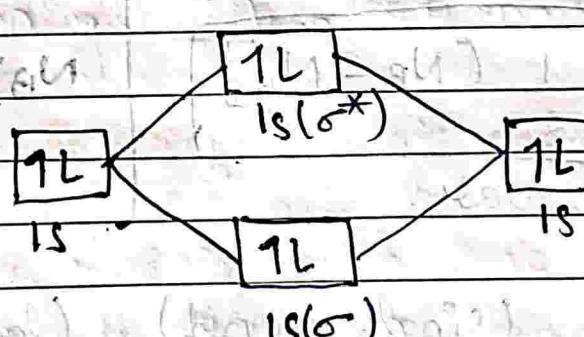
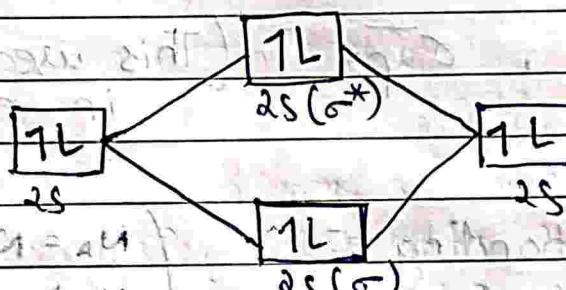
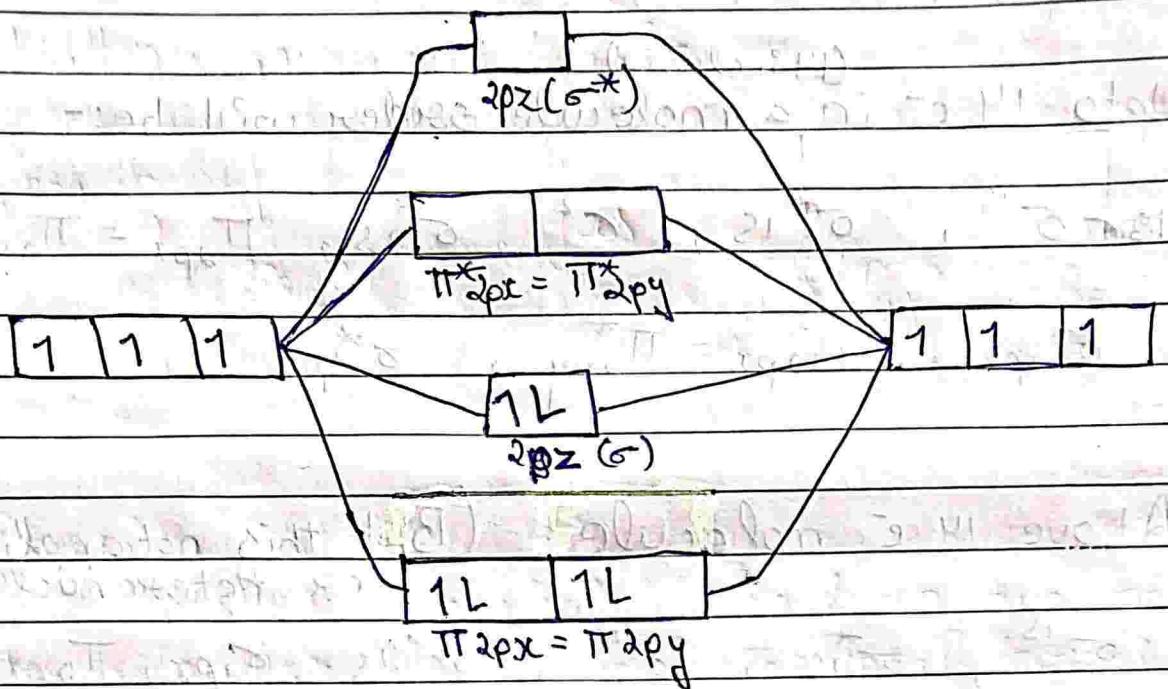
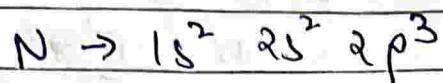
B.O = bond order.

 $B.O = 1$  (single bond) (least strength)

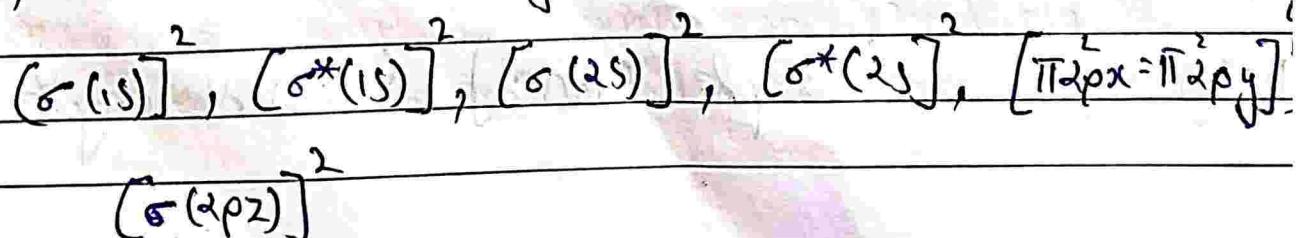
 $B.O = 2$  (Double bond)

 $B.O = 3$  (Triple Bond) (Most strength)

Example:- N<sub>2</sub>



So, its electronic configuration will be



## # Magnetic Moment ( $\mu$ )

$\mu = 0$  [The compound is non-magnetic in nature and is c/s Diamagnetic magnetic]

$\mu \neq 0$  [The compound is magnetic in nature and c/s Paramagnetic.]

Greater the  $\mu$ ; stronger will be the magnetic properties.

$$\mu = \sqrt{n(n+2)} \text{ B.M} \quad (\text{B.M - unit})$$

Bohr magneton

$n = \text{no. of unpaired } e^-$

## # Formula to calculate Hybridization-

$$H = \frac{1}{2} [nVe + M - c^+ + \bar{a}] \quad (\text{only take the magnitude of } c^+ \& \bar{a})$$

Ex:- ①  $H_3O^+$ , H of Oxygen

$$H = \frac{1}{2} [6 + 3 - 1 + 0] = \frac{8}{2} = 4 = Sp^3$$

②  $AlH_4^-$  (Al)

$$H = \frac{1}{2} [3 + 4 - 0 + 1] = \frac{8}{2} = 4 = Sp^3$$

Trick if No charge  $\rightarrow$  calculate no. of L.P & B.P & add.

# CHAPTER :- Periodic Classification of Elements

# Elements:- substances that are made up of only one type of atom.

There are 118 elements known till date, 94 are natural and 26 elements are man-made.

# Dobereiner:- "When the elements are arranged in the groups with similar properties are arranged in a increasing order of their atomic mass in a g.p. of elements called triplets, then the atomic mass of middle two elements will be the mean of other two elements."

$$\text{Ex:- } \begin{array}{c} \text{Li - 7} \\ \text{Na - 23} \\ \text{K - 39} \end{array} ] \quad \frac{7+39}{2} = \frac{46}{2} = 23 \text{ Na}$$

$$\text{Ca - 40} \quad \text{Cl - 35.5}$$

$$\text{Sr - 88} \quad \text{(1) Br - 81}$$

$$\text{Ba - 137} \quad \text{I - 127}$$

⇒ Drawbacks:-

It is only applicable only for few elements.

# Newland Octave Law:- "When elements are arranged in the increasing order of their A.M then the property of every 1<sup>st</sup> and 8<sup>th</sup> element is same."

1	2	3	4	5	6	7
H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca				

Drawbacks :-

- ① There is no mentioning of noble gas.
- ② It was not applicable for most elements having A.M more than A.M of Ca. (means for heavier elements).

# A-B-B de Chancourtoise :-

"In 1862 he arranged the elements in the increasing order of the A.M and placed them in a cylindrical table based upon their properties."

⇒ Drawback :- We do not study it because of its complex nature.

# Lothar Meyer :- He made a graphical representation of elements where he plotted a graph b/w Atomic volume Vs Atomic weight

Drawback :- Due to complexity it does not get much popularity

## # Mendeleev's Periodic Table :-

- There were 63 elements at that time.
- Always focused on properties of elements along with Empirical formulae of compounds formed by elements. (Hydrides and oxides).

Ex:-  $\text{CH}_4$ ,  $\text{CO}_2$

$\text{EH}_4$ ,  $\text{EO}_2 \rightarrow$  types

- He considered A.M of elements.
- He prepared cards denoting all the properties of an elements.

	$\text{EH}$	$\text{EH}_2$	$\text{EH}_3$	$\text{EH}_4$	$\text{EH}_3$	$\text{EH}_2$	$\text{EH}$
Series	I	II	III	IV	V	VI	VII
1	A/B	A/B	A/B	A/B	A/B	A/B	A/B
2	Li	Be	B	C	N	O	F
3	Na	Mg	Al	Si	P	S	Cl
4	K	Ca	Sc	Ti	V	Cr	Mn
5	Cu	Zn			As	Sc	Ba
	$\text{R}_2\text{O}$	$\text{RO}$	$\text{R}_2\text{O}_3$	$\text{RO}_2$	$\text{R}_2\text{O}_5$	$\text{RO}_3$	$\text{R}_2\text{O}_7$
							$\text{RQ}$

## # Mendeleev's Periodic Law :-

"Atomic mass is the periodic function of the properties of an elements."

#:- There are possibility of discovery of new elements in his table.

He predicted two elements :- Eka-Aluminium and Eka-Silica which now days are Gallium and germanium respectively.

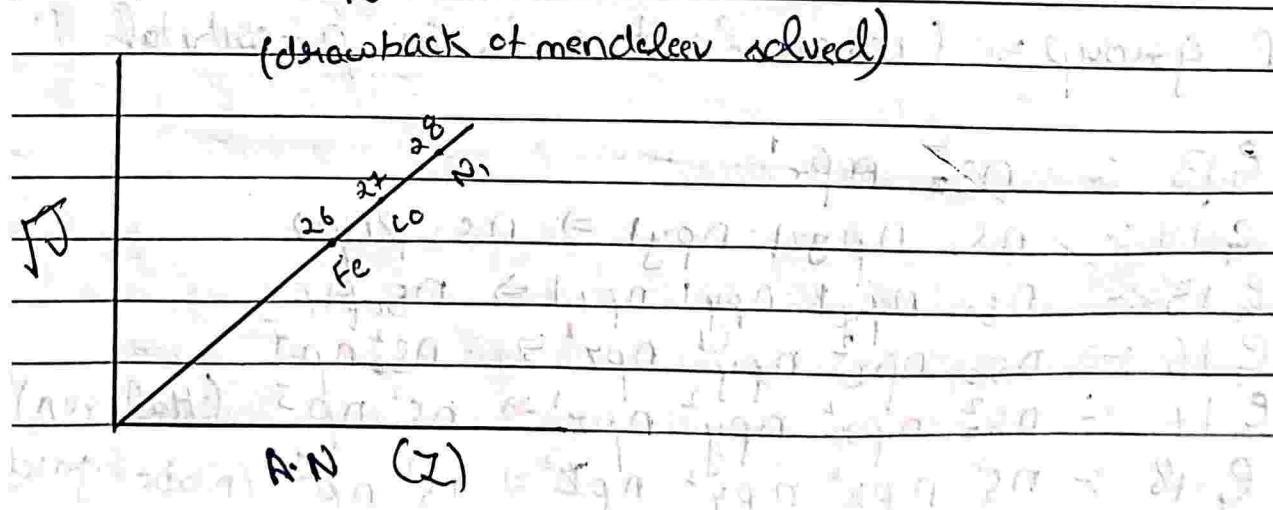
# There was no information of noble gases when this table is developed but this table was so flexible that after the discovery of noble gases a new group GROUP ZERO is created for them.

Drawbacks :-

1. Some elements mendeleev did not obey its own periodic law.  
Ex:- Fe Co Ni  
$$\begin{array}{ccc} 54 & 59 & \xrightarrow{\text{decreasing}} 58.5 \end{array}$$
  2. There is no place for isotopes in his periodic table.
  3. Position of hydrogen was not specified.

# # Modern Periodic Table

After the discovery of subatomic particles, a scientist Henry Moseley plotted a graph b/w Atomic Number and JD and then suggested that the elements should be arranged in the order of increasing Atomic Number.



# Modern Periodic Law by Henry Moseley :-  
 "The Physical and chemical properties of an element in the periodicity of its Atomic Number".

# Properties of every elements actually depends on their Atomic Number and electronic configuration.

\* Valency :- No. of e<sup>-</sup> present in the last shell or no. of e<sup>-</sup> required to complete the last shell is called valency.

\* There are 7 periods in MPT which moves from Left to Right and which are completely based on Principal quantum number.

# There are total 18 groups in MPT which moves up to down and which are based on subshells of valence shell namely s, p, d, f.

S-group. (when last e<sup>-</sup> is in s-orbital)

e.p-1 :- nS<sub>1</sub> Alkalimetals

e.p-2 :- nS<sub>2</sub> Alkali earth metals

P group. (when last e<sup>-</sup> is in p-orbital)

e.p-3 :- nS<sub>1</sub> nP<sub>3</sub>

e.p-4 :- nS<sub>2</sub> nP<sub>3</sub>1 nP<sub>3</sub>2  $\Rightarrow$  nS<sub>2</sub> nP<sub>6</sub>

e.p-5 :- nS<sub>2</sub> nP<sub>3</sub>1 nP<sub>3</sub>2 nP<sub>3</sub>3  $\Rightarrow$  nS<sub>2</sub> nP<sub>10</sub>

e.p-6 :- nS<sub>2</sub> nP<sub>3</sub>1 nP<sub>3</sub>2 nP<sub>3</sub>3 nP<sub>3</sub>4  $\Rightarrow$  nS<sub>2</sub> nP<sub>14</sub> (Halogen)

e.p-7 :- nS<sub>2</sub> nP<sub>3</sub>1 nP<sub>3</sub>2 nP<sub>3</sub>3 nP<sub>3</sub>4 nP<sub>3</sub>5  $\Rightarrow$  nS<sub>2</sub> nP<sub>18</sub> (Noble gases)

## # Exceptional Cases :-

### ① Position of Hydrogen :-

"Hydrogen shows similar properties to both G.P.1 alkali metals and G.P.17 Halogen and in some cases with G.P.14 Carbon family".

Properties Similar to .

G.P.1.

→ valence ~~shell~~ orbital :- ns<sup>1</sup>

→ only one e<sup>-</sup> in valence shell.

→ belongs to s-orbital.

G.P.17

→ require 1e<sup>-</sup> to complete its valence shell

→ it forms diatomic gas H<sub>2</sub>

# It shares e<sup>-</sup> to form bonds like P.P-14

"That's" why hydrogen is placed at the top of periodic table separately.

### ② Position of Helium :-

Helium's electronic configuration is similar to G.P-2 but it does not show similar properties like G.P.2, it shows inert properties hence it is placed among inert gases in G.P-18.

As its last shell is fully filled.

e

# Our periodic table can be divided in 4 parts based upon GPS:

→ S-block - G.1, G.2

→ P-block - G.13-G.18

D-block - G.3-G.12

f-block G.3

# #The S-Block Elements $n^2$

## GROUP 1.

- $ns^1$
- Alkali Metals
- Low Ionisation Enthalpy
- Most Reactive metals
- Li - Lithium
- Na - Sodium
- K - Potassium
- Rb - Rubidium
- Cs - Cesium
- Fr - Francium

→ It will donate  $1e^-$   
 → Li have covalent

## GROUP 2.

- $ns^2$
  - Alkali Earth metals.
  - Low Ionisation Enthalpy
  - Reactive metals.
  - Be - Beryllium
  - Mg - Magnesium
  - Ca - Calcium
  - Sr - Strontium
  - Ba - Barium
  - Ra - Radium
- It will donate  $2e^-$   
 → Be have covalent nature

# # The P-Block Element. ( $ns^2 np^1-6$ )

## G.P. 13. (Boron family)

~~B~~  $ns^2 np^1$  (Moving down metallic character increases)

## G.14 (Carbon family)

$ns^2 np^2$

## G.15 (Nitrogen family) Nitrogen - phosphorus

$ns^2 np^3$

## G.16 (Oxygen family) Chalcogens - arc forming

$ns^2 np^4$

## G.17 (Fluorine family) Halogens

$ns^2 np^5$

$\text{G}\cdot\text{I}7$  is highly reactive (only need  $1e^-$  to complete its octate).

## # THE D-BLOCK : - $ns^2(n-1)d^{1-10}$

They are also called as Transition metals

# These elements have  $(n-1)d$  orbital partially empty or filled because  $e^-$  first goes in  $4s$ .

Valence shell - 4<sup>th</sup> shell. (Ultimate orbital)

3<sup>rd</sup> shell (Penultimate orbital - 3d)

# These elements acts as bridge link b/w S/P block in a periodic table

→ All are metals

→ They are coloured compounds

→ They show magnetic character.

→ Variable valency.

Zn	→ Prede
Cd	→ T.E
Hg	becoz d-2 is fully fi

## # THE F-BLOCK $4f^{1-14} 5d^{0-1} 6s^2$

Inner transitional elements.

In lanthanum →

$4f \ 5d \ 6s^2$

as  $4f$  and  $5d$  are very close so they ignore  $1e^-$ 's position but when other  $e^-$  come they start following Aufbau's Rule and other  $e^-$  will go to  $4f$ . Hence the new block formed with inside the D-block.

Transition elements are those which have vacant d-orbital.

Page No. \_\_\_\_\_

Date : \_\_\_\_\_

$4f^{1-14} 5d^0 6s^2 \rightarrow$  Lanthanoid Series.

$5f^{1-14} 6d^0 7s^2 \rightarrow$  Actinoid Series.

# They are Radioactive

# Their half life is very small.

# IUPAC (International Union of Pure & Applied Chemistry).

→ Naming of elements with atomic numbers more than 100.

Digits	Name	Abbreviation.
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	quad	q
5	pent	p
6	hex	h
7	hept	g
8	oct	o
9	enn	e

Ex:- atomic number 189

1 :- un

Name :- unoctennium

8 :- oct

9 :- enn

" add ium at the end of the name"

Symbol :- Uoe

# The Elements in Periodic table are of 3 types :-

① Metals :- 78% of elements are metals.

→ Metals are e<sup>-</sup> donors.

→ They have high M.P & B.P.

→ They are solid at Room temperature  
exception :- Hg (mercury)

→ Good conductors of heat & electricity because of metallic bonding.

→ Malleability & Ductility.

→ Sonorous property.

② Non-Metals :- present on the top most right side of periodic table.

→ Non-metals are e<sup>-</sup> acceptors.

→ Non-metals are having low M.P and B.P.

→ They are either solid or gas at Room temperature.

→ Exception :- Br<sub>2</sub> (liquid)

→ Diamond (c) Highest M.P & B.P.

→ Bad conductivity.

③ Metalloids :-

These elements have properties of both metals and non-metals.

B

P-block

Si

Ge

As

Sb

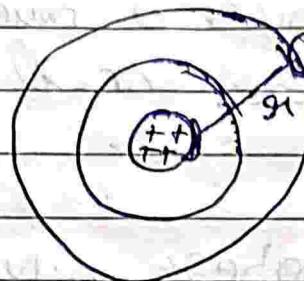
T

P O

# # Periodic Trends in the properties of Element.

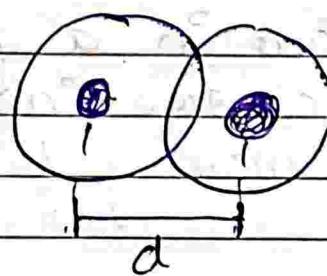
## (A) Physical Properties :-

(i) Atomic Radius :- Distance b/w the nucleus and the valence  $e^-$



### # Covalent Radius

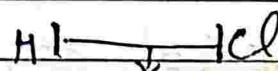
→ It is formed in non-metals.



$d$  = distance b/w  
two nucleus.

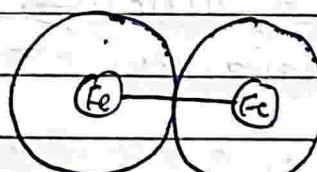
$$r_1 = \frac{d}{2}$$

if different atoms with  
different E.N.



Bond length  $\Rightarrow d = r_{1a} + r_{1b} - 0.9(Y_a - Y_b)$   
(Y = EN)

### # Metallic Radii

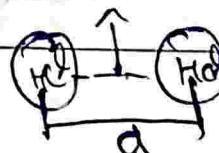
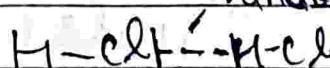


$d$  = ~~distance~~ The internuclear distance

$$r_1 = \frac{d}{2}$$

### # Van der Waal Radii

Vanderwaal dist. ~~dist. Radious~~ distance



$$r_1 = \frac{d}{2}$$

# On going Top - Bottom in a group; no. of shells increases.

# On going Left to Right in a period; no. of shells remain same.

# Effective Nuclear Charge

$\Rightarrow$  Power with which a nucleus hold / attract its valence shell e<sup>-</sup>. (ignoring inner shells)

$\Rightarrow$  As Atomic Number A.N increases no. of p increases and hence ~~because~~ nuclear charge increases.

# On moving top to bottom in a gp, as +pT,  $\Rightarrow$  Nuclear charge also increases but because of counter by screening effect "effective nuclear charge decreases"  
"In period left to Right effective nuclear charge T"

# Screening / Shielding Effect

$\Rightarrow$  The inner orbital e<sup>-</sup> repel to valence e<sup>-</sup> and stops or inhibits the nucleus to attract the valence e<sup>-</sup>.

S orbital > p orbital > d-orbital > f orbital

$\checkmark$   
Strongest  
Screening  
effect

discrete step  
negligible  
screening  
effect

## # Ionic Radius :-

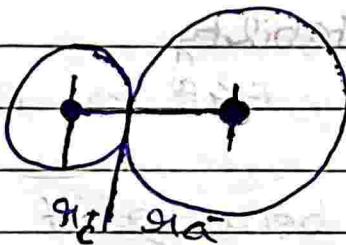
Atom  $\xrightarrow[\text{e}^-]{\text{loss}} \text{cation}^+$

Atom  $\xrightarrow[\text{e}^-]{\text{gain}} \text{Anion}^-$

Size of cation is atom is smaller than atom and size of anion is bigger than atom because of effective nuclear charge and interelectronic repulsions.

Size:- Cation < Atom < Anion

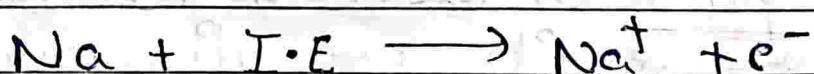
Ionic radius is internuclear distance b/w a cation and anion.



$$R_{\text{ionic}} = R_C + R_A^-$$

## # Ionisation Enthalpy

It is the energy supplied off given to one mole atom in gaseous form so that it releases one mole electrons.



Ionisation Enthalpy 1 is always smaller

\* In case of stable  $e^-$  configuration I.E is exceptionally high. (Jump 1 gp ahead)

Page No. \_\_\_\_\_ Date: \_\_\_\_\_

than Ionisation Enthalpy  $\Delta$ .

I.E also depend  
of subshell.

$$I.E_1 < I.E_2 < I.E_3$$

s > p > d > f

$$I.E \propto \frac{1}{\text{Atomic Radius}}$$

I.E  $\downarrow$   $\rightarrow$  I.E  $\uparrow$

## # Electron Gain Enthalpy :- (kJ/mol) (Electron affinity)

When an atom in gaseous form accepts an electron it releases the amount of Energy that is known as E.G.E.

Energy  $\propto$  Ease of taking  $e^-$   
released

\* Almost for all elements EGE is negative except Nobel gases

Be :- EGE is positive because of stable configuration.

More easily to gain  $e^-$  more negative EGE.

## # Electronegativity :-

It is the ability / power of atom to attract the shared pair of electron towards its side because of its effective nuclear charge in a covalent bond.

# It is not an experimental value (measurable).

Electronegativity is a comparative property.  
(we can only compare b/w two elements)

Page No.

Date :

## # Chemical Properties.

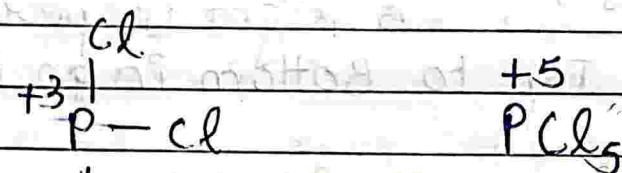
### 1. Valency or Oxidation State :-

No. of  $e^-$  present in valence shell

or 8 minus no. of  $e^-$  present in valence shell.

# Almost all d-f block atom and some of P-block atom shows various valency.

→ that why we say them oxidation state as the charge created on an atom after forming bonds till the associated atoms are removed.



### 2. Diagonal Relationships

Every first element of each gp shows similar properties as that of the second element of its next corresponding group. (This is because charge/size ratio)

### 3. Typical Elements

→ Other all the elements of the gp, shows the similar properties to these elements & typical elements are (period-3) element.

# # TRENDS

## 1. Atomic Radius

- (i) Moving Left to Right in period it decreases
- (ii) Moving Top to Bottom in gp it increases

## 2. Effective Nuclear charge

- (i) Moving Left to Right in period it increases
- (ii) Moving Top to Bottom in gp it also decreases

## 3. Shielding effect

- (i) Moving Left to Right in period it decreases
- (ii) Moving Top to Bottom in gp it increases

## 4. Ionisation Enthalpy

- (i) Moving Left to Right in period it increases
- (ii) Moving Top to Bottom in gp it decreases

## 5. Electronegativity

- (i) Moving Left to Right in period it increases
- (ii) Moving Top to Bottom in gp it decreases

## 6. Electron Gain Enthalpy

(i) It becomes more negative

(ii) It becomes less negative on moving  
left to right in a period and Top  
to bottom in a group

(iii) It becomes more negative on moving  
from left to Right in a period

# # Exceptions / Dependency

① opposite to the trend atomic Radius of oxygen is more than that of Nitrogen.

$$O > N$$

$$\text{also } Ni < Cu < Zn$$

This is stable configuration of Cu, Zn & Nitrogen  
As atomic radius depends on the no. of shells, Effective nuclear charge and Intersi e<sup>-</sup> repulsion

because of stable configuration N, Zn, Cu have less intersi e<sup>-</sup> repulsions.

② Effective Nuclear charge depends on number of protons vs number of electrons.

③ Shielding Effect depends on Atomic Radius, and Number of shells b/w Nucleus and valence shell and intersi e<sup>-</sup> repulsions.

④ Ionisation Energy of Be > B ; N > O  
it depends on stability of electronic configuration, atomic Radius, effective nuclear charge.

Because Be & N electronic configuration is stable they require more energy to remove an electron.

as well as, Mg > Al &  $\begin{cases} Ca > Ga \\ Ca \approx Ga \end{cases}$

as well as P > S, As > Se

③ The Electron gain Enthalpy depends upon two things :- Inter-electronic repulsion and Effective nuclear charge.

~~- EGE due to interelectronic repulsion~~

~~- EGE due to effective nuclear charge~~

- EGE  $\propto$  Effective Nuclear charge

- EGE  $\propto$    
 Inter-e<sup>-</sup> repulsion

Fluorine (F) & oxygen (O) have less negative EGE enthalpy than (Cl) and sulphur (S) even though respectively even they have more Effective Nuclear charge,

this is because of their small size they have high inter-electronic repulsion, so it is not easy to put an e<sup>-</sup> in them than Cl & S which have high ENC & less IER.

# Second EGE for an atom is always positive:

Oxygen has less<sup>^n</sup> EGE than S, Se, Te  
(inter-e<sup>-</sup> repulsion)

most -ve  $\rightarrow$  S, Se, Te, O

N have +ve EGE ; it increases down the group N till Sb & Bi have least -ve.

# # Pauling Scale for Electronegativity.

A = 1<sup>st</sup> atom

B = 2<sup>nd</sup> atom

$x_A$  = Electronegativity of A

$x_B$  = Electronegativity of B

This scale is used for comparison of Electronegativity of different atoms.

$E_d$  = dissociation Energy.

$$|x_A - x_B| = \sqrt{E_d(AB) - \frac{E_d(AA) + E_d(BB)}{2}} \times (ev)^{-\frac{1}{2}}$$

## \* Lattice Energy

$$L.E \propto \frac{q_1 q_2}{r}$$

$q_i \rightarrow$  literally mag. of ch.

