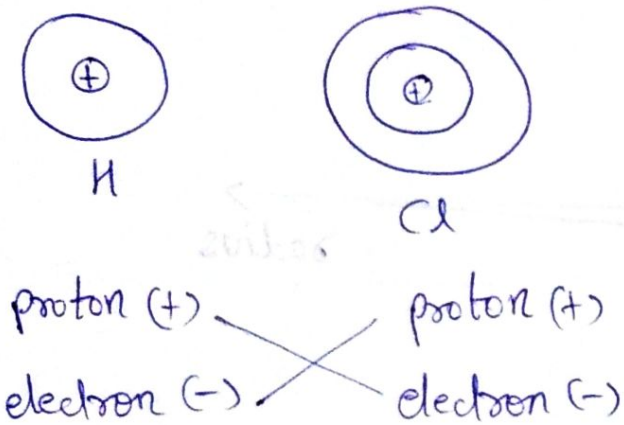
Chemical Bonding

2018 → 3 questions, 12 marks

Why Chemical Bonding?

1. To gain stability or gain nearest inest gas configuration.
2. To lower its potential energy of the system.

How chemical Bonding?



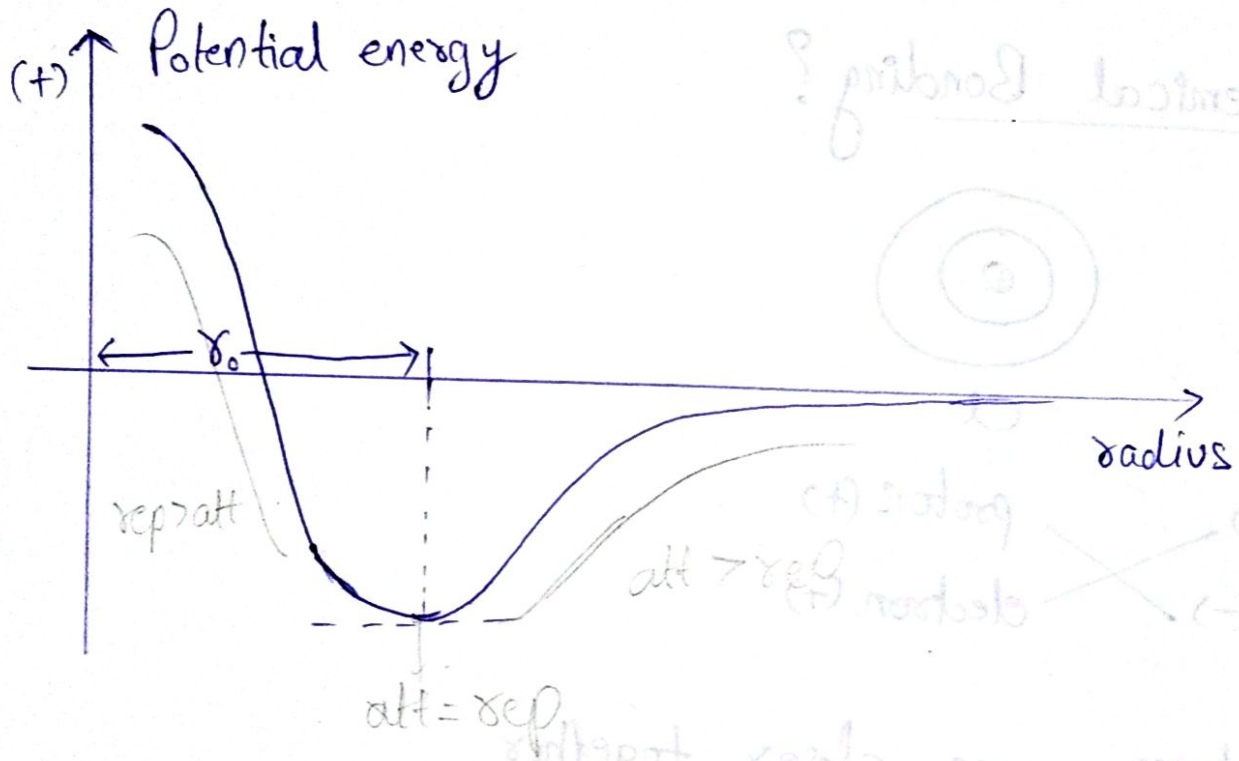
Why atoms come closer together

- Attraction b/w nucleus of hydrogen & electrons of chlorine
- Attraction b/w nucleus of chlorine and electrons of hydrogen

Lower potential energy of the system

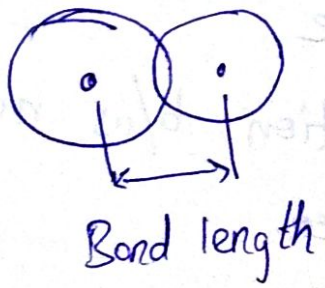
$$P.E = K\frac{Q_1 Q_2}{r}$$

$r\uparrow \Rightarrow P.E\downarrow$ or $r\downarrow \Rightarrow P.E\uparrow$



Bond length

r_0 is called bond-length



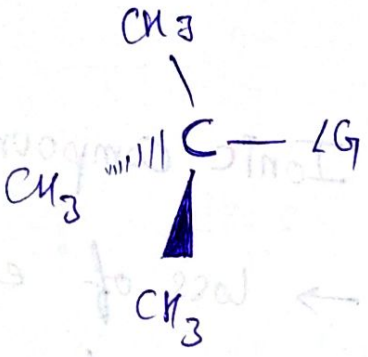
$r = r_0$ ($F_{attraction} = F_{repulsion}$)

$r > r_0$ ($F_{attraction} > F_{repulsion}$)

$r < r_0$ ($F_{attraction} < F_{repulsion}$)

Intersecting Facts

- If we bring two atoms more closer to its bond length then one atom goes out of plane. (steric hinderance)



Types of Bond

Strong Bond

1. Ionic Bond
2. Covalent Bond
3. Co-ordinate Bond
4. Metallic Bond

(200 – 400 kJ/mol)
Bond Energy

Weak Bond

1. Vander wall force
2. Hydrogen Bond

(2 – 40 kJ/mol)
Bond energy

Ionic Bond

- Bond formed between cation & anion.
- Bond formed by complete transfer of electrons.

Ionic compound formation

Only 3 steps, by which all Ionic compounds are formed:

1. formation of cation → loss of e^-
2. formation of anion → gain of e^-
3. formation of Lattice. → combine cation & anion

Eg. Show the formation of $NaCl_{(s)}$:

{loss or gain ⇓ must be gases state}

→ 1) formation of cation

$$Na_{(g)} \longrightarrow Na^+_{(g)} + e^-$$

2) formation of anion

$$Cl_{(g)} + e^- \longrightarrow Cl^{\ominus}_{(g)}$$

3) formation of Lattice

$$Na^{\oplus}_{(g)} + Cl^{\ominus}_{(g)} \longrightarrow NaCl_{(s)}$$

g. Show the formation of $MgO_{(s)}$

→ 1) formation of cation.

$$Mg_{(g)} \longrightarrow Mg^{2\oplus}_{(g)} + 2e^-$$

2) formation of anion

$$O_{(g)} + 2e^- \longrightarrow O^{2-}_{(g)}$$

3) formation of Lattice

$$Mg^{2\oplus}_{(g)} + O^{2-}_{(g)} \longrightarrow MgO_{(s)}$$

<u>Statement type Questions</u>

1. Ionic Bond is always formed between different nature elements.

2. One should be cation and other should be anion

OR

One should be metal and other should be non-metal

OR

One should have low I.E and other should have high electron affinity.

Energy Involved in Ionic Bond

1. formation of cation → loss of electron
   (+) (Ionization energy)
2. formation of anion → gain of electron
   (−) (Electron Affinity)
3. formation of Lattice → Lattice Energy (−)

(+) → Endothermic
(−) → Exothermic

Lattice Energy:

- Energy released when 1 mole of ionic compound is formed from isolated gaseous cation & anion

Eg. formation of $HCl_{(s)}$

1. Cation formation ---- Ionization Energy:

$$H_{(g)} \longrightarrow H^+_{(g)} + e^- \quad \text{(energy absorb)}$$

2. Anion formation ... Electron Affinity:

$$Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)} \quad \text{(energy release)}$$

3. Lattice Energy

$$H^+_{(g)} + Cl^{\ominus}_{(g)} \longrightarrow HCl_{(s)} \quad \text{(Energy release)}$$

Comparison of Ionic character:

- Ionic character $\propto \Delta E.N.$

More the electro negativity difference $\Rightarrow$ more will be ionic character in the bond

How: easily cation $\rightarrow$ Cs – F $\rightarrow$ easily anion

0.6 4

$= 3.4$

Q. Arrange in order of ↑ ionic character

(a) LiF, NaF, KF, RbF, CsF

$\rightarrow$ I.C $\propto \Delta E.N$

LiF < NaF < KF < RbF < CsF

$$\begin{cases} LiF = 4 - 1 = 3 \\ NaF = 4 - 0.9 = 3.1 \\ KF = 4 - 0.8 = 3.2 \\ RbF = 4 - 0.7 = 3.3 \\ CsF = 4 - 0.6 = 3.4 \end{cases}$$

b) NaCl, NaBr, NaI

c) MgF_2, $MgCl_2$, $MgBr_2$, MgI_2

Determination of Lattice Energy
(Born - Haber cycle)

Diff b/w Heat of formation (ΔH_f) & lattice energy (L.E)

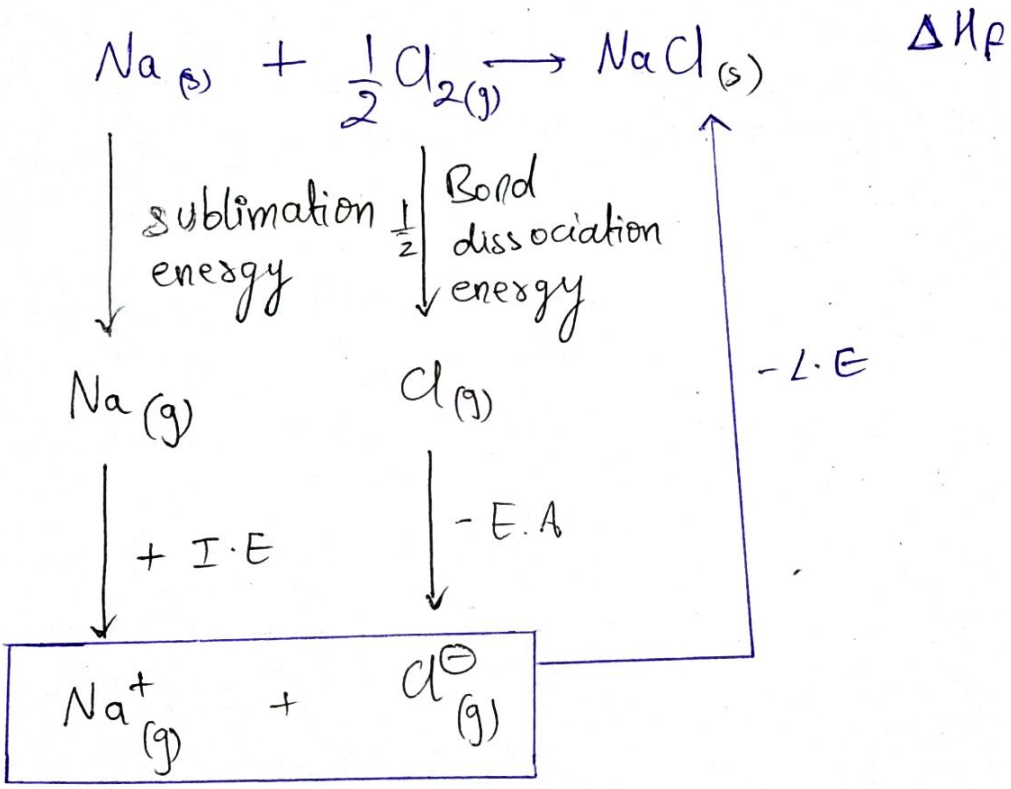
ΔH_f	L.E
Energy change when one mole of compound is formed from its elements which is in pure & stable form	Energy released when one mole of compound is formed from isolated gaseous ions
eg $Na_{(s)} + \frac{1}{2}Cl_{2(g)} \longrightarrow NaCl_{(s)}$ ↑ Stable form of Na ↑ Stable form of Cl	eg $H^{\oplus}_{(g)} + Cl^{\ominus}_{(g)} \longrightarrow HCl_{(s)}$

Q) Draw born-Haber cycle of NaCl

OR

find L.E in the formation of $NaCl_{(s)}$

→ Write ΔH_f of $NaCl_{(s)}$



$$\Delta H_f = \text{Sublimation} + I.E + \frac{B.D.E}{2} - E.A - L.E$$

Q. Draw Born-Haber cycle of MgO.

OR

Find L.E. in the formation of $MgO_{(s)}$

→ $Mg_{(s)} + \frac{1}{2} O_{2\,(g)} \longrightarrow MgO_{(s)}$ ΔH_f

Factors Affecting Lattice - Energy

(i) Charge on ions

(ii) Size of ions

$$\boxed{L.E \propto \frac{\text{Charge}}{\text{Size}}}$$

Between the charge & size ; charge is dominating factor

Q) Compare Lattice Energy

i) $\overset{+1}{Na}\overset{-1}{F}$ $\quad MgF_2$ $\quad AlF_3$

→ $NaF < MgF_2 < AlF_3$

ii) $\overset{+1}{Na_2}O$ $\quad MgO$ $\quad Al_2O_3$

→ $Na_2O < MgO < Al_2O_3$

Q. Compare Lattice Energy

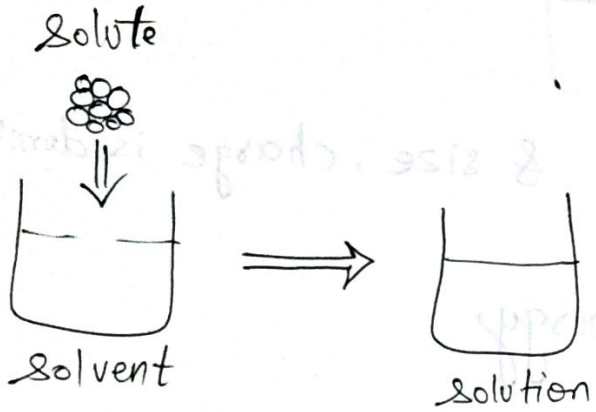
$\overset{+1}{Na}\overset{-1}{Cl} < \overset{+2}{Mg}\overset{-2}{O}$

$$\boxed{L.E \propto \frac{\text{charge on ion}}{\text{Size of ion}}}$$

Lecture- 2

Solubility

- Solubility is the process of surrounding solute with solvent. (solvation)



Ionic Bond & Solubility

- Ionic compounds kha soluble honge?

OR

Ionic compounds kis me soluble honge?

Like dissolves Like

- Ionic compounds are polar in nature, so ionic compounds are soluble in polar solid (solvent).

Hydration

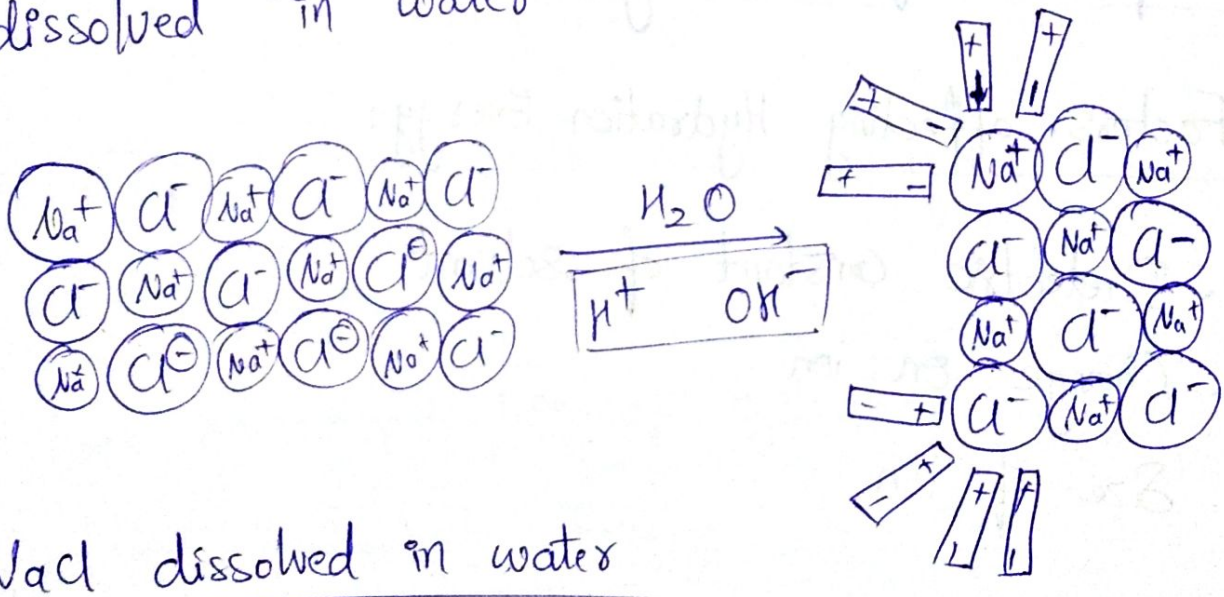
- when solvent is water, same process is called Hydration.
- Hydration is the process of surrounding solute with water.

What about energy involved?

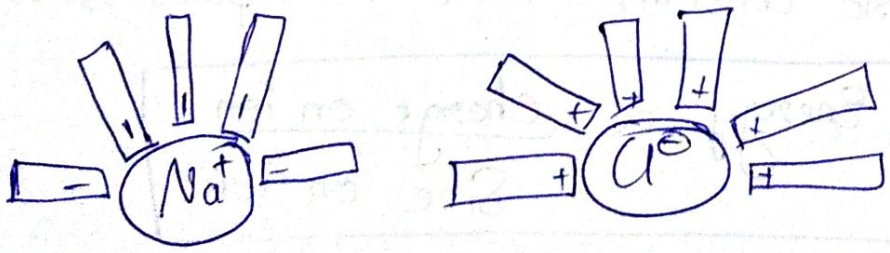
→ In solubility, solvation energy is released. (EXO)

In hydration, hydration energy is released (EXO)

What actually happens when ionic compound is dissolved in water

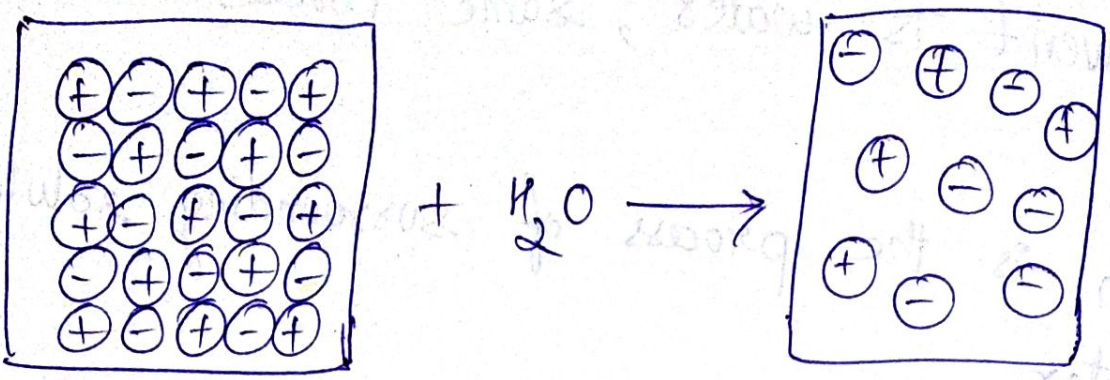


NaCl dissolved in water



$Na^+_{(aq)}$ + $Cl^-_{(aq)}$

when Ionic compound dissolves in H_2O



$Na^+_{(aq)}$ & $Cl^-_{(aq)}$

Cations and anions alag alag ho jate hain

Type 2: Questions

Questions based on hydration Energy

Comparison of Solubility / H.E

- Factors affecting Hydration Energy:

1. Dielectric constant of solvent
2. Charge on ion
3. Size of ion

1. More dielectric constant means more polar solvent
2. $\boxed{\text{Hydration Energy} \propto \frac{\text{Charge on ion}}{\text{Size on ion}}}$

eg. Compare the solubility

F^-, O^{2-}, N^{3-}

→ $\boxed{\text{Solubility} \propto \text{Hydration Energy} \propto \text{charge of ion}}$

$F^{-1} < O^{-2} < N^{-3}$

eg. Compare hydration Energy.

$Li^+, Na^+, K^+, Rb^+, Cs^+$

⇒ $Li^+, Na^+, K^+, Rb^+, Cs^+$

$\xrightarrow{\text{Size} \uparrow}$

$\xrightarrow{\text{hydration-}E \downarrow}$

$\therefore Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$

Type 3: Questions

Questions on Application of Hydration Energy

Application of Hydration Energy

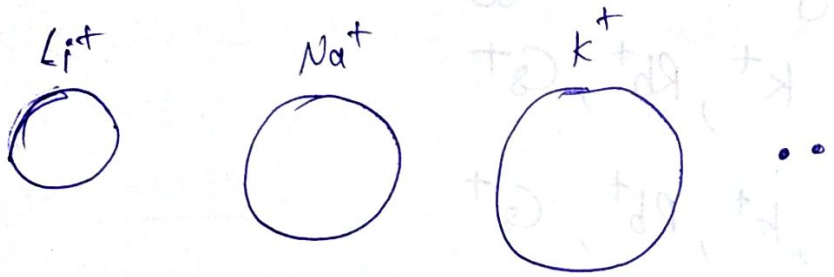
i) Size of Hydrated ion

ii) Mobility of ion

iii) Conductivity of ion

i) Size of Hydrated ion

- Smaller size of gaseous ion means, higher size of hydrated ion.
- Size of gaseous ion is small ⇒ Charge density is more



- Size of gaseous ions increases, density of charge decreases

↓ means ↓

Charge density decreases ⇒ less water molecules attract

Eg. $Li^+_{(g)} < Na^+_{(g)} < K^+_{(g)} < Rb^+_{(g)}$

Size ↑

$Li^+_{(aq)} > Na^+_{(aq)} > K^+_{(aq)} > Rb^+_{(aq)}$

Size of ion in aqueous sol^n

ii) Mobility of ion

$$\text{Mobility} \propto \frac{1}{\text{size in aqueous solution}}$$

iii) Conductivity of ion

$$\text{Conductivity} \propto \frac{1}{\text{size in aqueous sol}^n}$$

Q. which of the following has least mobility & conductivity

A) $Be^{+2}_{(aq)}$ ✓ B) $Sr^{+2}_{(aq)}$ C) $Ca^{+2}_{(aq)}$

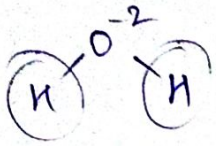
D) All have same mobility

⟹ Be, Mg, Ca, Sr, Ba (size ↑)

$$\text{mobility \& conductivity} \propto \frac{1}{\text{size in aq. sol}^n}$$

Eg*. F^- & Na^+ have same charge & almost same size. In which case F^- or Na^+ H.E. released is more? or same!

⟹ H^+ OH^- F^- ; H^+ OH^- Na^+



In Flourine there will be interaction between two hydrogen & in sodium there will be tween one hydrogen

∴ Jaha pe interaction jagdya waha energy jadya release hoga.

∴ $F^- > Na^+$

Type 4 : Questions

Questions in which both H.E & L.E is involved.

Factors involved in solubility :

i) Lattice energy (L.E)

ii) Hydration energy (H.E.)

i) Lattice energy (L.E) : Energy released when 1 mole of compound is formed from its isolated gaseous cations & anions.

ii) Hydration energy (H.E.) : Energy released when 1 mole of compound completely breaks down into cations & anions

$|H.E| > |L.E| \Rightarrow$ Compound soluble

$|H.E| < |L.E| \Rightarrow$ Compound insoluble

Eg. $BeSO_4$ is soluble, while $BaSO_4$ not? why?

$\Rightarrow$ In $BeSO_4 \rightarrow |H.E| > |L.E|$

In $BaSO_4 \rightarrow |H.E| < |L.E|$

Eg. which is more soluble LiF or LiCl?

$\Rightarrow$ In Both Li^+ is common, so deciding factor is F & Cl.

$$\boxed{H.E \propto \frac{\text{charge on ion}}{\text{size of ion}}}$$

Since, charge is same $\therefore$ Size $\uparrow \Rightarrow H.E \downarrow$

Size of F < Size of Cl.

$\therefore \boxed{H.E \text{ of } LiF > H.E \text{ of } LiCl}$

$\therefore$ LiF is more soluble than LiCl

Eg. which is more soluble $NaClO_4$ or $KClO_4$?

→ In $NaClO_4$ & $KClO_4$, ClO_4 is common.

∴ Deciding Factor is Na & K.

$$\boxed{H.E \propto \frac{\text{Charge on ion}}{\text{Size of ion}}}$$

Since, charge is same. So $H.E \propto \frac{1}{\text{Size of ion}}$

Size of K > Size of Na

∴ $(H.E)_K < (H.E)_{Na}$

H
Li
Na
K
Rb

Size ⇒ ↑

$$\boxed{\therefore NaClO_4 > KClO_4}$$

Eg. Which is least soluble in H_2O.

A) LiF B) NaF C) KF D) RbF

~~Eg. which is least soluble in H_2O~~

~~A) LiF B)~~

⇒ Since, F is common. So it has nothing to do.

which Has more H.E, it will be more soluble.

$$\boxed{(H.E) \propto \frac{\text{charge on ion}}{\text{Size of ion}}}$$

Since, Charge is same, so. $H.E \propto \frac{1}{\text{Size of ion}}$

H Li Na K Rb Cs Fr

Size ↑

∴ Size of Rb > K > Na > Li

∴ H.E of Li > Na > K > Rb

∴ Solble of LiF > NaF > KF > RbF

$\boxed{\therefore \text{ RbF is least soloble.}}$

Eg. Arrange in order of solubility –

BeF_2, $BeCl_2$, $BeBr_2$, BeI_2

→ Since, Be is common. So it has nothing to do. which has more H.E, it will be more soluble

$$\boxed{H.E \propto \frac{\text{charge on ion}}{\text{Size of ion}}}$$

Since, charge is same, so. $\boxed{H.E \propto \frac{1}{\text{Size of ion}}}$

F Cl Br I

Size ↑

∴ H.E of F > Cl > Br > I

∴ Solubility order: $\boxed{BeF_2 > BeCl_2 > BeBr_2 > BeI_2}$

VVIP

Lecture 3

FAJAN'S Rule

What is Fajan's Rule?

With the help of FAJAN's rule, we can find covalent character in Ionic Bond.

- Covalent character in Ionic bond.

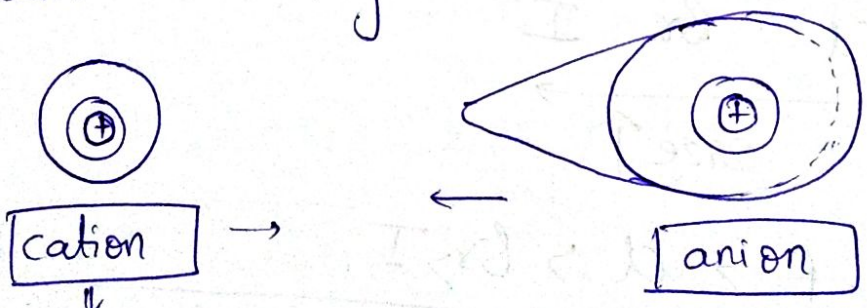
Aisa bhi hota hai ???

Ionic bond me ionic character hona chahiye.

Covalent character in Ionic BonD

1. Neither bond is 100% ionic nor 100% covalent
2. Each bond has some ionic character into it & some covalent character into it.

Diagram se Fajan's Rule samjhte hain



high Zeff

So, attracts electron cloud of anion

when cation & anion approaches each other to form ionic bond.

then

Due to high Z_{eff} of cation, it starts attracting electron cloud of anion due to which sharing of electron cloud takes place and hence covalent character develops.

As a result of this attraction of electron cloud,

↓

Sharing of electron cloud takes place

↓

Sharing means covalent

↓

Hence, covalent character develops in Ionic Bond

Imp. terms in Fajan's Rule

- Polarizing Power (cation)
- Polarization (Anion)

Polarizing Power ⟶ Polarize karne ki Power

Polarization ⟶ Polarize Hone ki Power

Definition:

- Polarizing Power (Cation +) : Power of cation to deform (attract) electron cloud of anion is called polarizing power.

- Polarization (anion -) : Deformation of electron cloud of anion by cation is called polarization.

Factors affecting Polarizing Power & Polarization

Polarizing Power

- Polarizing power of cation is due to Z_{eff}
- Polarizing Power $\propto Z_{eff} \propto \frac{\text{charge of ion}}{\text{Size of ion}}$

P.N.G.C → Pseudo Noble gas Configuration.

- Cation having Pseudo noble gas Configuration have more polarizing power as compared to cation having noble gas configuration.

Noble gas configuration → $8e^-$ in outermost orbit
Pseudo noble gas configuration → $18e^-$ in outermost orbit

eg. Noble gas Configuration

$Na^+ \rightarrow 1s^2\ 2s^2\ 2p^6$ → 8e

$Ca^{+2} \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ → 8e

$K^+ \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ → 8e

5s 5p 5d 5f
4s 4p 4d 4f
3s 3p 3d
2s 2p
1s

eg. Pseudo noble gas configuration.

$Cu^{+} \rightarrow 1s^2\ 2s^2\ 2p^6\ \underline{3s^2}\ \underline{3p^6}\ \underline{3d^{10}} \rightarrow 18e^-$

$Zn^{+2} \rightarrow 1s^2\ 2s^2\ 2p^6\ \underline{3s^2}\ \underline{3p^6}\ \underline{3d^{10}} \rightarrow 18e^-$

⇒ d-Block elements show P.N.G.C

Polarization

- Polarization $\propto$ charge
- Polarization $\propto$ size

Factors Supporting covalent character Increment in a Bond

1. High polarizing power of cation
2. High polarization of anion

⇒ more covalent character

TRICK FOR Fajan's RULE

S C L A < PNGC

(+) (−)

↓

smaller cation Larger anion

Eg. Compare polarizing power of cation?

Li^+, Na^+, K^+, Rb^+

→ SCLA → tho jayda polarizing power

SC → Li^+

∴ $Li^+ > Na^+ > K^+ > Rb^+$

Eg. Compare polarizing power of cation?

Na^+, Cu^+

→ Cu^+ → P.N.G.C

∴ $Cu^+ > Na^+$

Eg. Compare polarization of anion.

i) N^-, N^{-2}, N^{-3}

→ SCLA → tho jayda polarization.

$\therefore N^{-3} > N^{-2} > N^-$

ii) F^-, Cl^-, Br^-, I^-

→ SCLA → tho jayda polarization

$\therefore I^- > Br^- > Cl^- > F^-$

iii) NO_2^-, NO_3^-

→ SCLA → tho jayda polarization

$\boxed{NO_3^- > NO_2^-}$ → oxygen jayda hai tho larger anion.

iv) SO_3^{-2}, SO_4^{-2}

→ SCLA → tho jayda polarization

$SO_4^{-2} > SO_3^{-2}$ → oxygen jayda hai tho larger anion

Applications of FAJAN's rule:

- Polarizing power $\propto$ Covalent character $\propto$ Polarization

- Covalent character $\propto \frac{1}{M.P} \propto \frac{1}{B.P} \propto \frac{1}{\text{Solubility}}$

Eg. Arrange in increasing order of covalent character.
LiBr, NaBr, KBr, RbBr

→ Anion is common (Br). So it has nothing to do.

SCLA → jayda Covalent character

∴ LiBr > NaBr > KBr > RbBr

Eg. Arrange in increasing order of covalent character.
Na_2O, Al_2O_3, MgO

→ $\overset{+}{Na}\,\overset{-2}{O}$ $\quad \overset{+3}{Al}\,\overset{-2}{O}$ $\quad \overset{+2}{Mg}\,\overset{-2}{O}$

O^{-2} is common.

SCLA → jayda covalent character

(+)↑ ⇒ Size↓

∴ $Na_2O < MgO < Al_2O_3$

Eg. Arrange in increasing order of covalent character.

NaCl , CuCl

→ ∵ CuCl → P.N.G.C (Pseudo Noble gas configuration)

∴ CuCl > NaCl

Eg. Arrange in M.P order :

i) $SnCl_2$, $SnCl_4$

→ SCLA ∝ covalent character joyda ∝ $\frac{1}{M.P}$

$Sn^{+2}\ Cl^{-1}$ $Sn^{+4}\ Cl^{-1}$

(+)↑ ⇒ Size↓

∴ Covalent character : $SnCl_4 > SnCl_2$

∴ M.P : $SnCl_2 > SnCl_4$

ii) $PbCl_2$, $PbCl_4$

$Pb^{+2}\ Cl^{-1}$ $Pb^{+4}\ Cl^{-1}$

(+)↑ ⇒ Size↓

Covalent character : $PbCl_4 > PbCl_2$

∴ M.P : $PbCl_2 > PbCl_4$

aracter.

nfiguration

Eg. Arrange in solubility order

BeF_2 , $BeCl_2$, $BeBr_2$, BeI_2

→ $Be^{+2}F^{-1}$ $Be^{+2}Cl^{-}$ $Be^{+2}Br^{-1}$ $Be^{+2}I^{-1}$

SCLA ⇒ Covalent character jayda.

Covalent order : $BeI_2 > BeBr_2 > BeCl_2 > BeF_2$

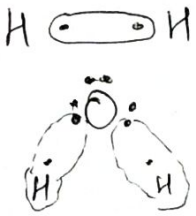
$$\boxed{\text{Solubility} \propto \frac{1}{\text{Covalent Charater}}}$$

∴ solubility : $BeI_2 < BeBr_2 < BeCl_2 < BeF_2$

COVALENT BOND

- Bond formed by sharing of electrons are called covalent Bond.
- Covalent bond is formed between <u>Non-metal</u> & <u>Non-metal</u>

Difference b/w Ionic & Covalent Bond.

Ionic Bond	Covalent Bond
→ Formed by complete transfer of electrons	→ formed by sharing of electron
→ Chlorine atom steal one electron from sodium $Na^+ \; Cl^-$	
Ionic bond is formed between metal & non-metal	Covalent Bond is formed between non-metal & non-metal

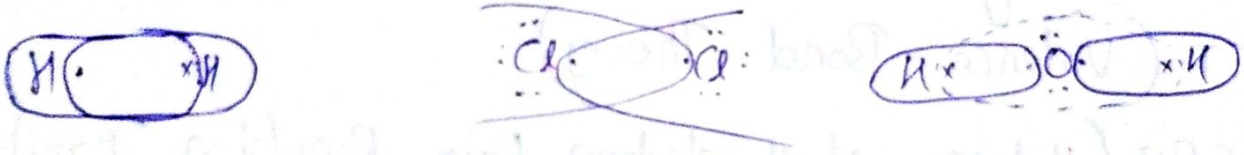
Types of covalent Bond

1. Single Covalent Bond → sharing of 2 electrons.
2. Double Covalent Bond → sharing of 4 electrons.
3. Triple Covalent Bond → sharing of 6 electrons.

Single Bond

- Sharing of two electrons

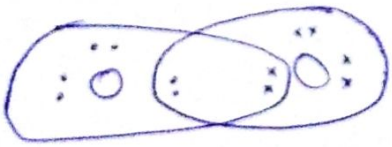
eg. H_2 Cl_2 H_2O



Double Bond

- Sharing of four electrons

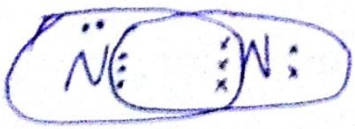
eg O_2 →



Triple Bond

- sharing of 6 electrons

eg. N_2 →



Condition for formation of Covalent Bond.

- The electro negativity difference between two atoms either zero or less.

  e.g. H_2, N_2, P_2, Cl_2, F_2 PCl_5, PCl_3, H_2SO_4

$\Delta E.N \propto$ Ionic character

Theories of Covalent Bond

1. Lewis theory
2. VBT (Valence Bond Theory)
3. VSEPR (Valence shell electron Pair Repulsion Theory)
4. MOT (Molecular orbital Theory)

Lewis Theory

Lecture - 4

According to this theory

- Valence shell electrons are shared between atoms in order to complete its octet.
- (For Hydrogen its duet)

Lewis Dot structure

Steps to draw LDS

1. Find central atom
2. Find total valence electron
3. Arrange atoms symmetrically around central atom.
4. Arrange remaining electrons.

Identify Central Atom

1) Atom having least electro negativity
2) Atom which is least in number
3) Atom having highest atomic mass
4) Atom having largest size
5) Atom which will form multiple bonds & have Vacant orbital.

Steps to draw LDS

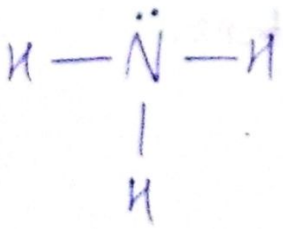
1. Central atom
2. Arrange remaining atoms symmetrically around Central atom.

Number of bonds formed to satisfy octet

C	N	O	F	P	S	Cl
4	3	2	1	3-5	2-6	1-7

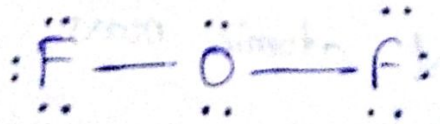
Eg. Draw LDS

1) NH_3

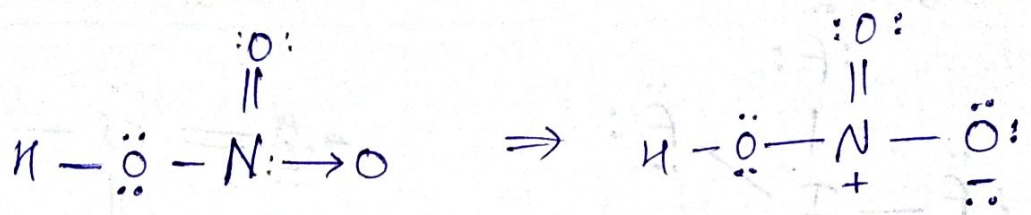


1 bond → 3 lone pair
2 bond → 2 lone pair
3 bond → 1 lone pair

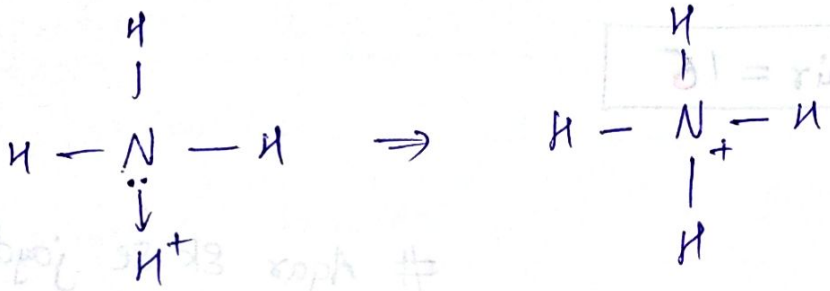
2) OF_2



3) HNO_3



4) NH_4^+



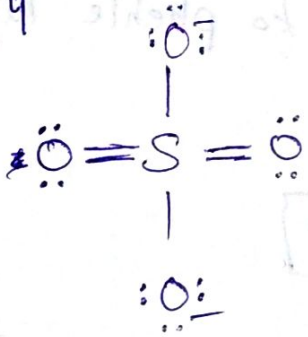
SKB

Jabhi (+) charge ayee → tho Central metal atom pe dal dena!

&

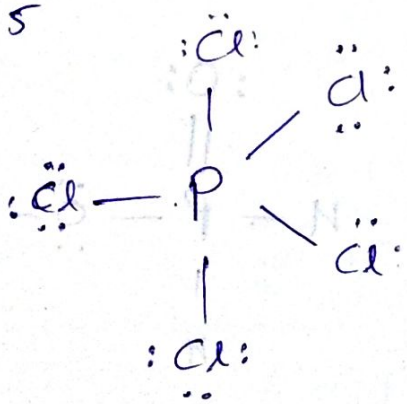
(→) → Sorrounding atom pe dal dena.

5) SO_4^{-2}



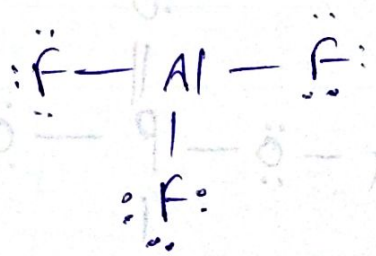
Total Lone pair = 10

6) PCl_5

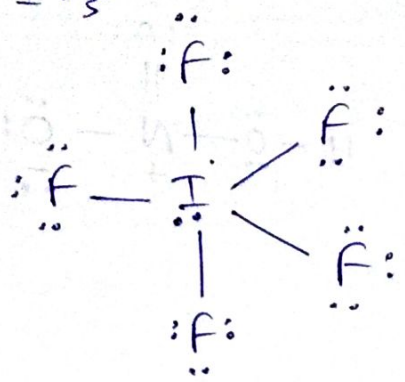


Total lone pair = 15

7) AlF_3

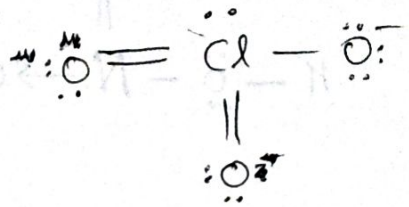


7) IF_5

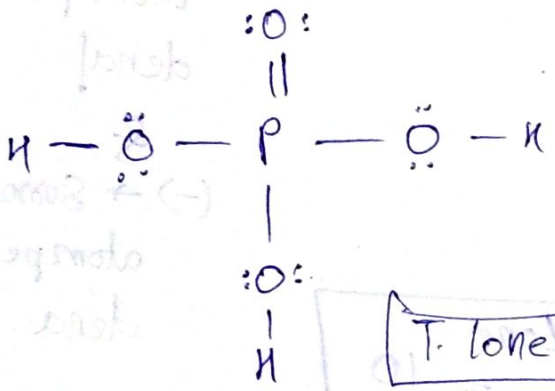


Total lone pair = 16

11) ClO_3^-



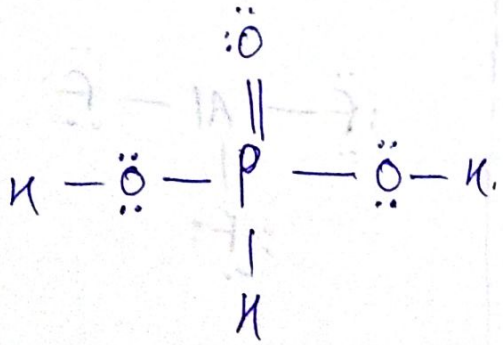
8) H_3PO_4



T. lone pair = 8

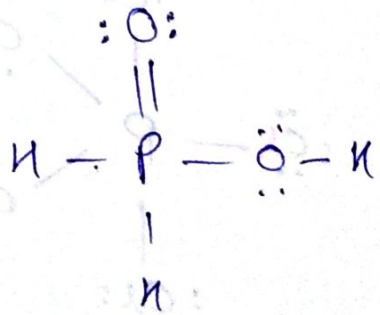
Agar ek se jyada surrounding atom hai tho mose E.N atom ko phehle lagna!

9) H_3PO_3



Total lone pair = 6

10) H_3PO_2



Total lone pair = 4

DRAWBACK of LEWIS THEORY

i) It could not explain shape of any compound.

ii) It did not tell anything about bond length, bond order, bond energy etc.

Important drawback

A/c to lewis

Compounds having less than $8e^-$

OR

having more than $8e^-$ after sharing should not exist.

Lewis theory failed? why?

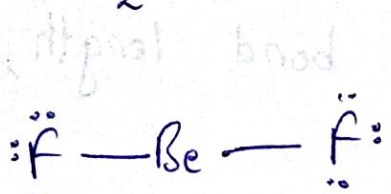
- Compounds having less electrons than octet are called hypo valent
- Compounds having more electrons than octet are called hypervalent

Octet bole tho 8 electron

In Nature hypovalent, hypervalent exists so Lewis theory failed.

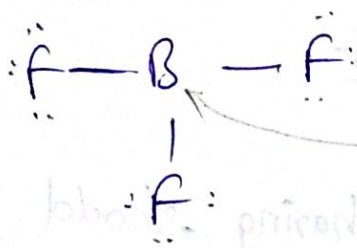
Eg. of Hypovalent :

i) BeF_2



Hypovalent
(only $4e^-$)
Octet not complete

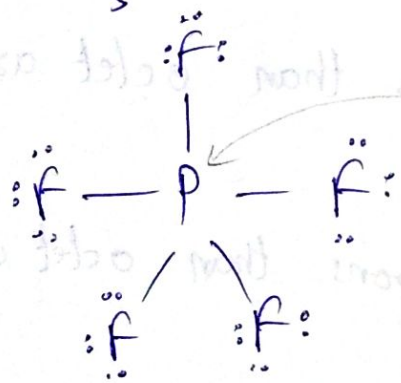
ii) BF_3



Hypovalent
(only $6e^-$)

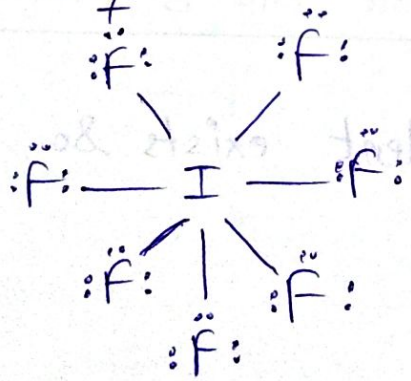
Eg. of Hypervalent :

i) PF_5



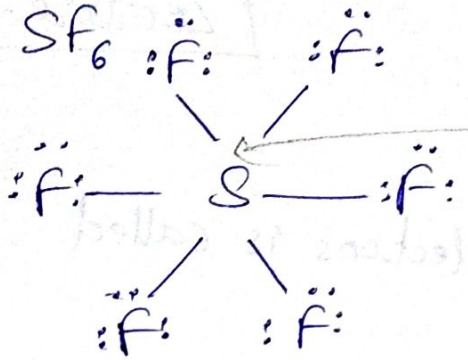
more than $8e^-$
($10e^-$)
(Hypervalent)

ii) IF_7



more than $8e^-$
$14e^-$
(Hypervalent)

iii) SF_6



More than $8e^-$

$(12e^-)$

(hypervalent)

RESONANCE

Lecture - 5

Que 1 : What is Resonance?

→ Delocalization of π (pi) electrons is called Resonance.

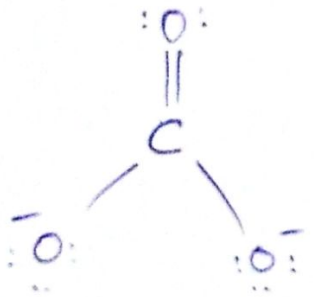
Que 2 : π (pi) electron kon se electrons ko bolte hain?

→ Electrons involved in π bond formation.

Eg. CH=CH , O=O

Que 3 : Why Resonance?

Let's understand this with the help of ex. of CO_3^{2-}

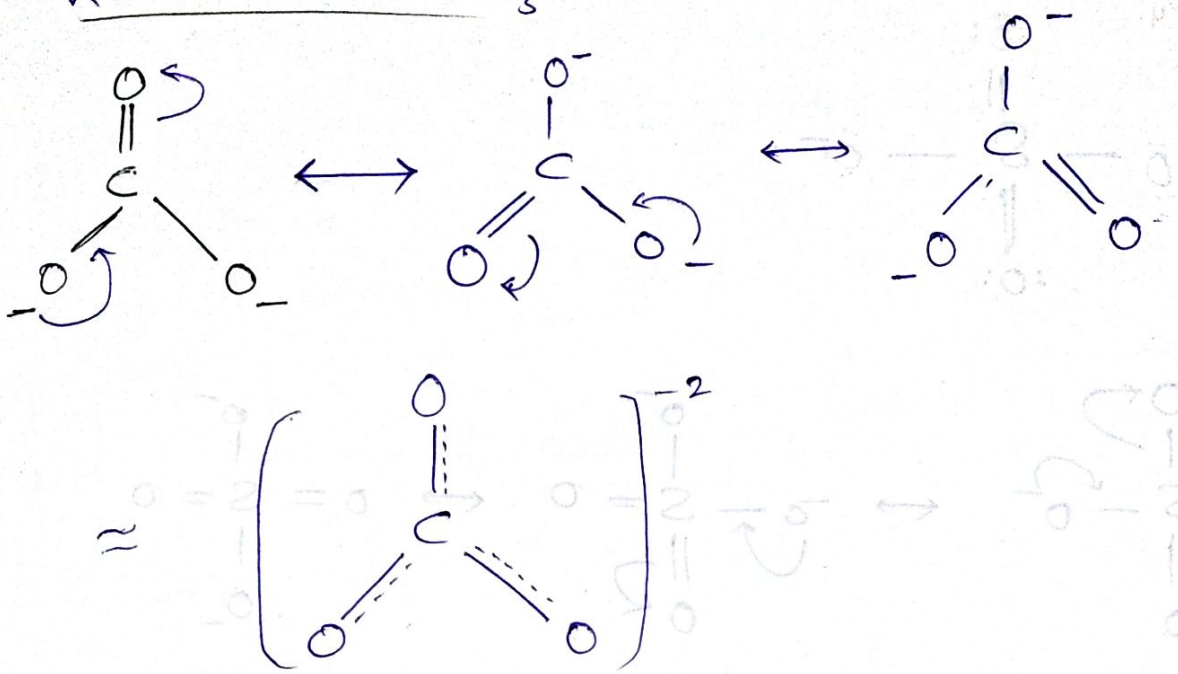


one double bond
&
two single bond

- Theoretically double bond should have less bond length than single bond.
- Experimentally we found that all the C–O bonds have equal bond length

Explained by Resonance.

Resonance in CO_3^{2-}



8. where Resonance?

- Resonance occurs where there is Conjugation

Conjugation

i) C = C — C = C

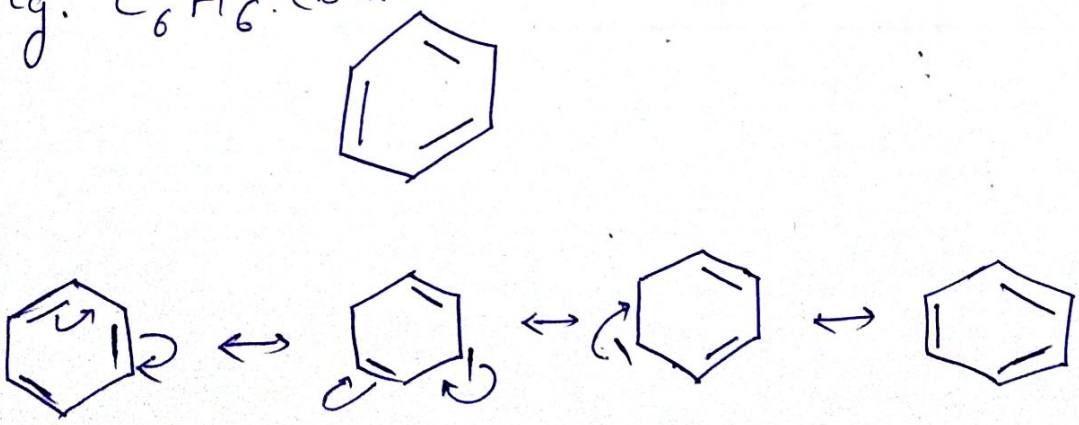
ii) C = C — X⊖

iii) C = C — X:⊖

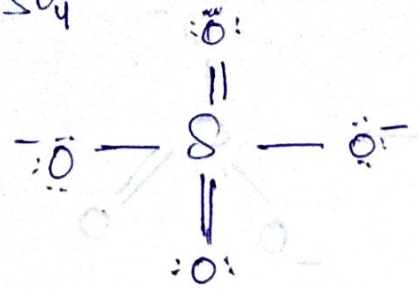
(iv) C = C — X⊙

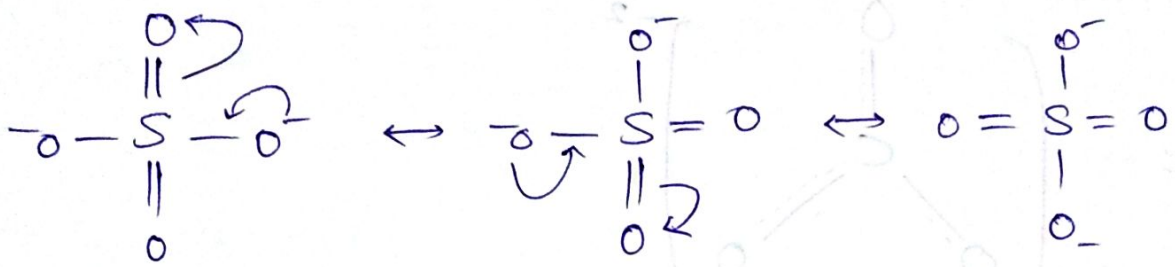
v) C = C — X⊕

Eg. C_6H_6 : (Benzene)

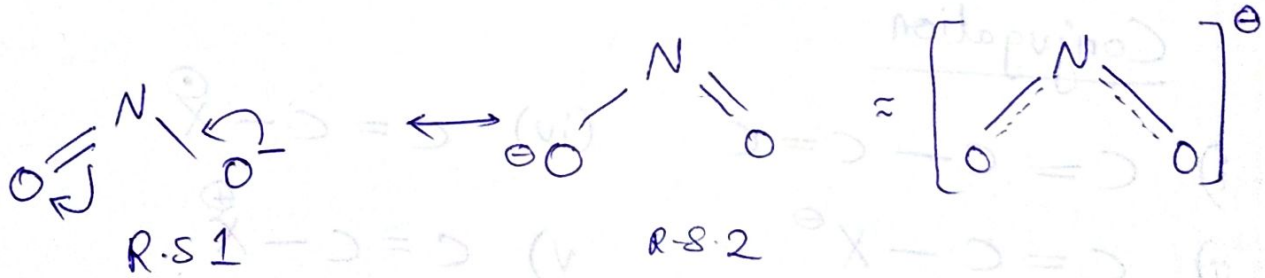


Eg. SO_4^{-2}





Eg. $NO_2^{\ominus}$



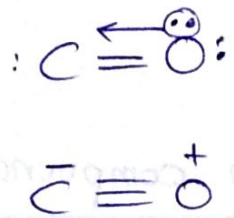
Q. why delocalization of only π e^- occurs?

BOND ORDER

- Single Bond → Bond order → 1
- Double Bond → Bond order → 2
- Triple Bond → Bond order → 3

eg. Calculate Bond order.

i) CO



$\bar{C} \equiv \overset{+}{O}$

∴ Bond order = 3

ii) CO_2

Bond order = 2

TRICK FOR CALCULATING B.O in radicals

$$\text{Bond order} = \text{Valency of surrounding atom} + \frac{\text{Charge on compound}}{\text{number of surrounding atom}}$$

Eg. 1) PO_4^{-3}

$$\text{Bond order} = \text{Valency of surrounding atom} + \frac{\text{Charge on compound}}{\text{Number of surrounding atom}}$$

$$B.O = 2 + \frac{(-3)}{4} = \frac{8-3}{4} = \frac{5}{4} = 1.25$$

$$= 1.25$$

2) ClO_4^-

$$B.O = \text{Valency of surrounding atom} + \frac{\text{Charge on compound}}{\text{Number of surrounding atom}}$$

$$= 2 + \frac{(-1)}{4} = \frac{8-1}{4} = \frac{7}{4} = 1.75$$

$$= 1.75$$

3) SO_4^{-2}

$$B.O = \text{Valency of surrounding atom} + \frac{\text{Charge on compound}}{\text{Number of surrounding atom}}$$

$$= 2 + \frac{(-2)}{4} = \frac{8-2}{4} = \frac{\cancel{6}^{3}}{\cancel{4}_{2}} = \frac{3}{2}$$

$$= 1.5$$

P.Y.Q

Q. Compa

CO

→ i) C

B.O

ii) CO_2

B.

iii)

B

P.Y.Q

Q. Compare Bond order

CO $\quad$ CO_2 $\quad$ CO_3^{-2}

→ i) CO → $\bar{C} \equiv \overset{+}{O}$

B.O = 3

ii) CO_2 → $O = C = O$

B.O = 2

iii) CO_3^{-2}

$$B.O = \text{Valency of surrounding atom} + \frac{\text{Charge on compound}}{\text{Number of surrounding atom}}$$

$$= 2 + \frac{(-2)}{3} = \frac{6-2}{3} = \frac{4}{3} = 1.33$$

$$\boxed{\therefore CO > CO_2 > CO_3^{-2}}$$

Eg. Calculate B.O of ClO_4^-, ClO_3^-, ClO_2^-

i) ClO_4^-

$\rightarrow$ B.O = Valency of surrounding atom + $\frac{\text{Charge on Compound}}{\text{Surrounding atoms}}$

$= 2 + \frac{(-1)}{4} = \frac{8-1}{4} = \frac{7}{4} = 1.75$

ii) ClO_3^-

$\rightarrow$ B.O $= 2 + \frac{(-1)}{3} = \frac{6-1}{3} = \frac{5}{3} = 1.67$

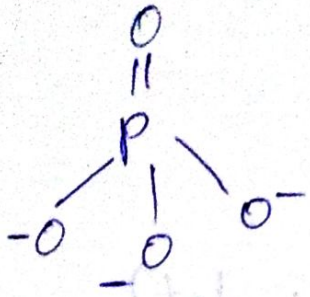
iii) ClO_2^-

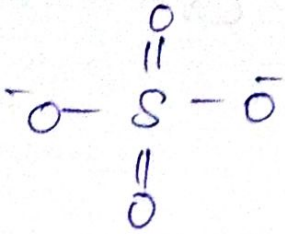
$\rightarrow$ B.O $= 2 + \frac{(-1)}{2} = 2 - 0.5 = 1.5$

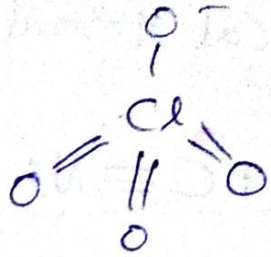
$\boxed{\therefore ClO_4^- > ClO_3^- > ClO_2^-}$

IMP Information:

PO_4^{-3}	SO_4^{-2}	ClO_4^-
$\downarrow$	$\downarrow$	$\downarrow$
1.25	1.5	1.75
		$\downarrow$
More single Bond	equal no. of single & double bond	More double bond







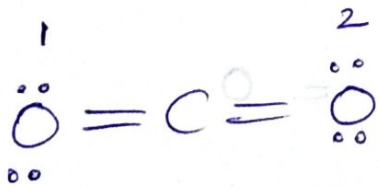
$$\text{Bond order} \propto \text{bond strength} \propto \text{Bond energy} \propto \text{stability} \propto \frac{1}{\text{Bond length}}$$

FORMAL Charge

- To calculate charge on individual atom. Concept of formal charge is used.

$$F.C = \text{Valence electron} - (\text{Bond} + \text{dots})$$

Eg. CO_2



$$(F.C)_{oxygen_1} = 6 - (2+4) = 6-6 = 0$$

$$(F.C)_{carbon} = 4 - (4+0) = 0$$

$$(F.C)_{oxygen_2} = 0$$

Eg. Cal formal charge CN^-

$:\overset{\ominus}{C} \equiv N:$

$(F.C)_{\text{~~Charge~~ Carbon}}$ = Valence e^- − (Bond + dot)

$= 4 - (3 + 2)$

$= 4 - 5 = -1$

$(F.C)_{\text{Nitrogen}}$ = valence e^- − (Bond + dot)

$= 5 - (3 + 2) = 5 - 5 = 0$

Eg. Calculate formal charge : OCN^-

→ $\ddot{N} \equiv C - \ddot{O}:^-$

$(F.C)_N = 5 - (3 + 2) = 5 - 5 = 0$

$(F.C)_C = 4 - (4 + 0) = 0$

$(F.C)_O = 6 - (1 + 6) = 6 - 7 = -1$

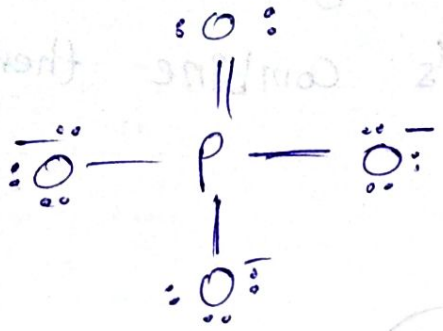
Eg. Calculate formal charge

PO_4^{-3}

$_{15}P$ → 2, 8, 5

$_{16}S$ → 2, 8, 6

→



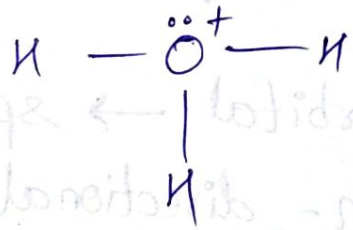
$(F.C)_P = 5 - (5 + 0) = 0$

$(F.C)_{O_{\parallel}} = 6 - (2 + 4) = 6 - 6 = 0$

$(F.C)_{O^-} = 6 - (1 + 6) = 6 - 7 = -1$

Eg. SO_4^{-2}

Eg. H_3O^+



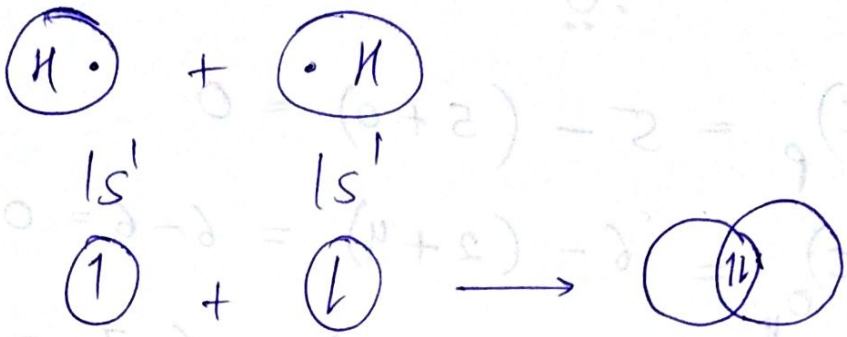
$(F.C)_O = 6 - (3 + 2)$

$= 6 - 5 = +1$

Valence Bond Theory

Lecture - 6

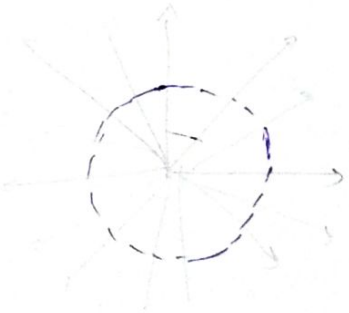
- When atomic orbital's combine then covalent bond is formed.



We are saying that orbitals will combine
Aao orbitals ki khabar lein ab!!

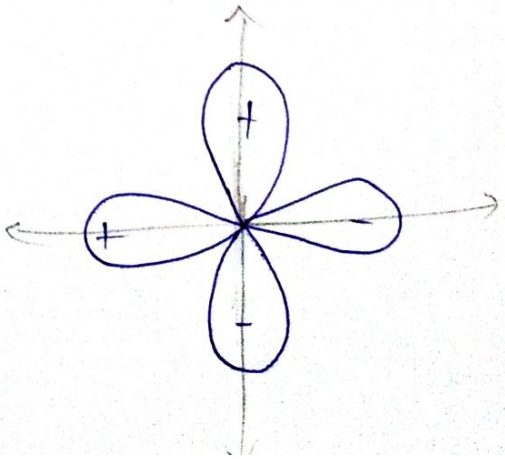
Shapes of orbitals:

'S' orbital → spherical
(non-directional)

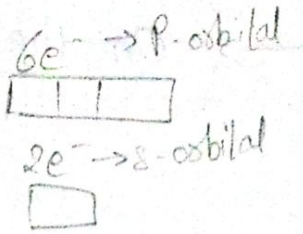


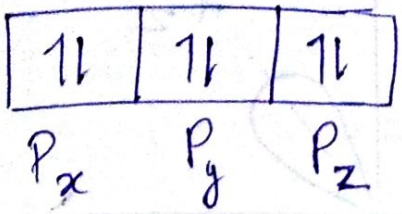
'P' orbital → dumbell
(directional)

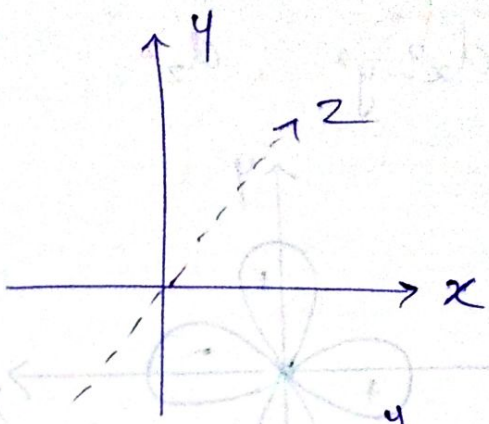
'd' orbital → double dumbell
(directional)

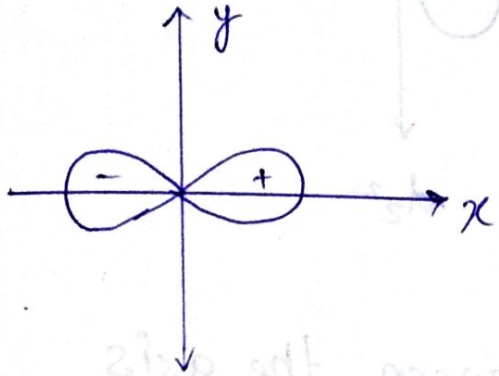


P- orbital

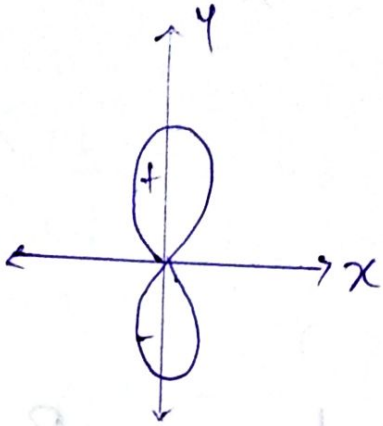




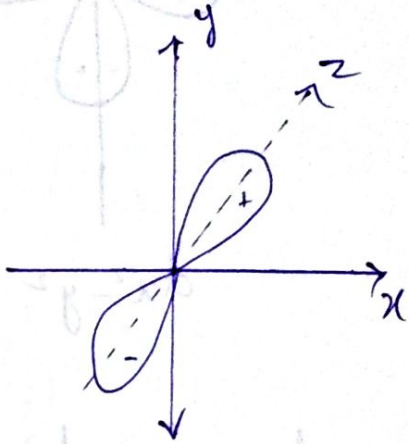




P_x

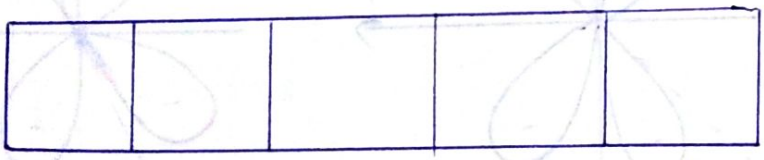


P_y



P_z

"D" orbital

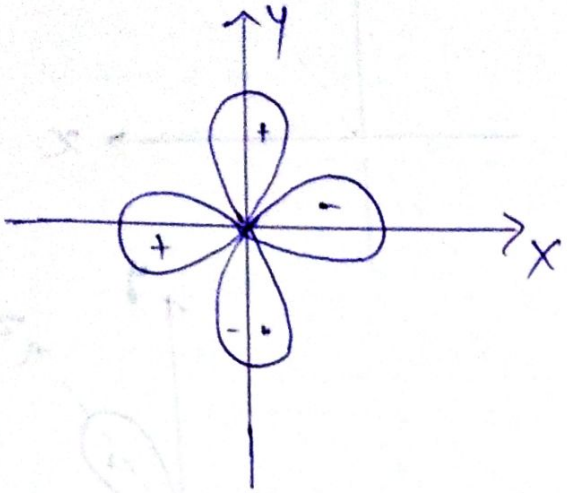


d_{xy} d_{yz} d_{zx} $d_{x^2-y^2}$ d_{z^2}

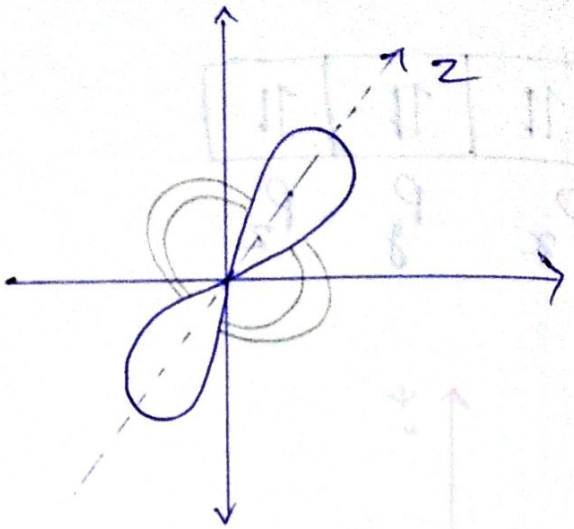
d_{xy}, d_{yz}, d_{zx} → Between the axis

$d_{x^2-y^2}$, d_{z^2} → On the axis

$d_{x^2-y^2}$, d_{z^2} → <u>On the axis</u>

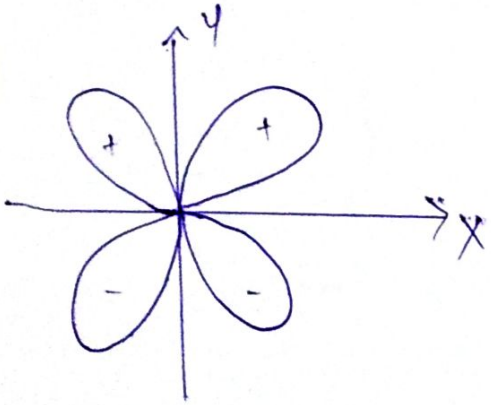


$d_{x^2-y^2}$

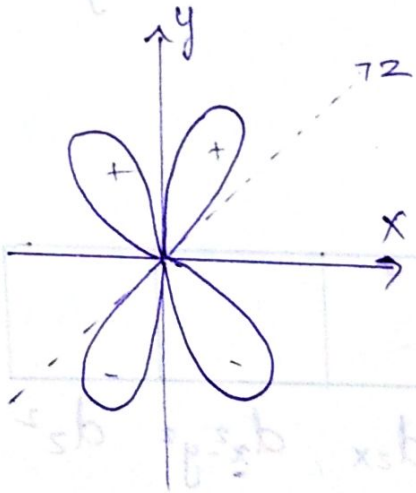


d_{z^2}

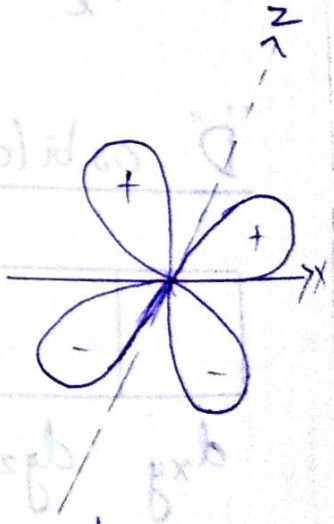
d_{xy} , d_{yz} , d_{zx} → <u>Between the axis</u>



d_{xy}



d_{yz}



d_{zx}

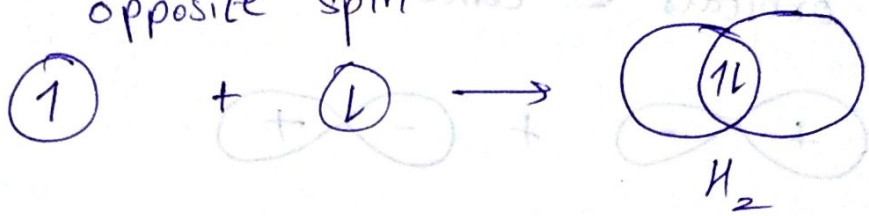
Important points of VBT :

1. Overlapping orbitals must be half filled & both should have opposite spin.

eg.

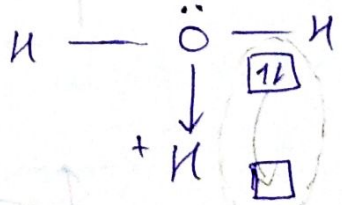
H →	← H
1s¹	1s¹
[↑]	[↓]

half filled
opposite spin



2. In case of Co-ordinate bond empty orbital of one atom overlaps with fully filled orbital of another atom.

Eg. H_3O^+ :



3. More overlapping of orbitals means more strong bond means more stability means more bond energy means less bond length

Extent of overlapping $\propto$ strength of Bond $\propto$ stability $\propto$ Bond energy

$\propto \frac{1}{\text{Bond length}}$

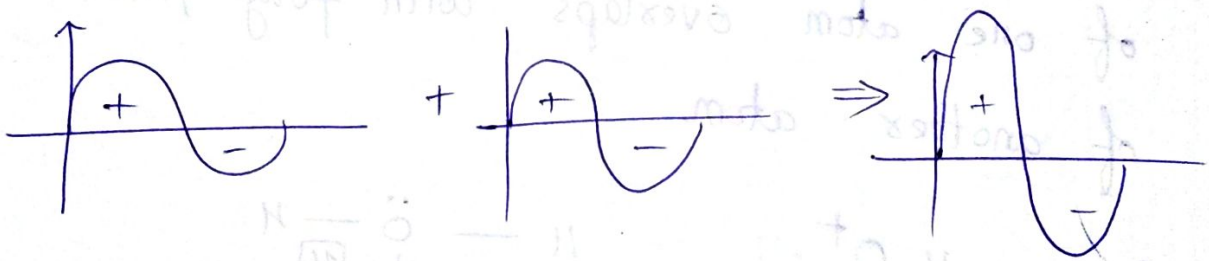
(4) Overlapping lobes should have same sign.

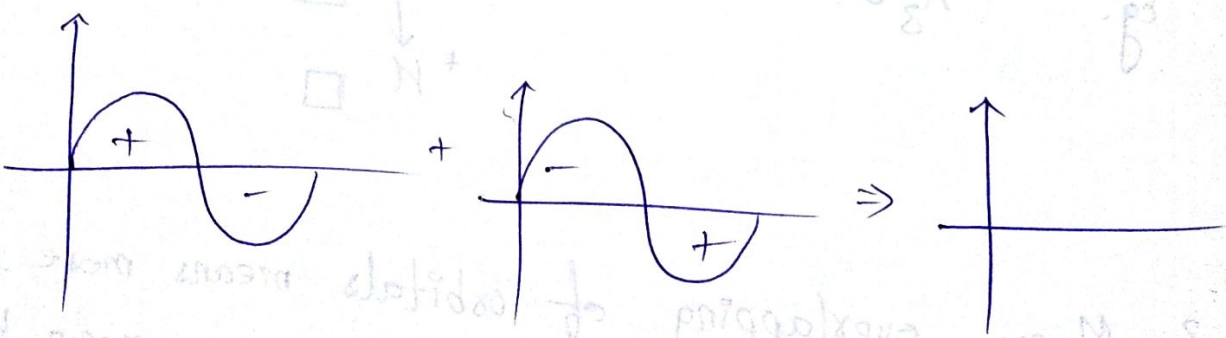
→

orbitals = Combine

Same sign

Another method





(5) More closer to nucleus means more extent of overlapping.

- more bond strong → 1-1 > 1-3 > 2-3

(6) In case of orbitals, more directional means more overlapping.

- more strong bond → s-s < s-p < p-p

Overlapping

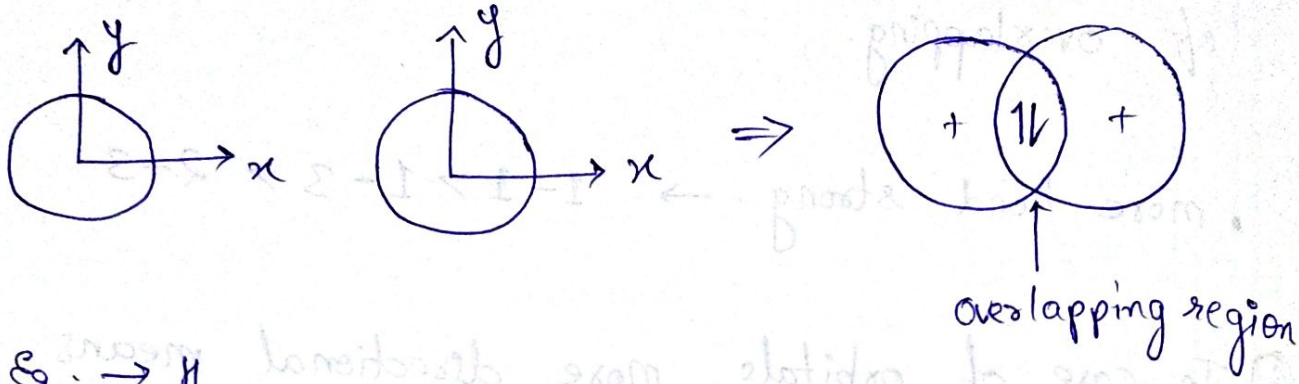
1. Axial overlap or Head-on overlap
2. Side-ways overlap or Parallel overlap.

Axial overlapping - (Head on Overlap)

- Overlapping of orbital along inter-nuclear axis is called axial-overlapping.

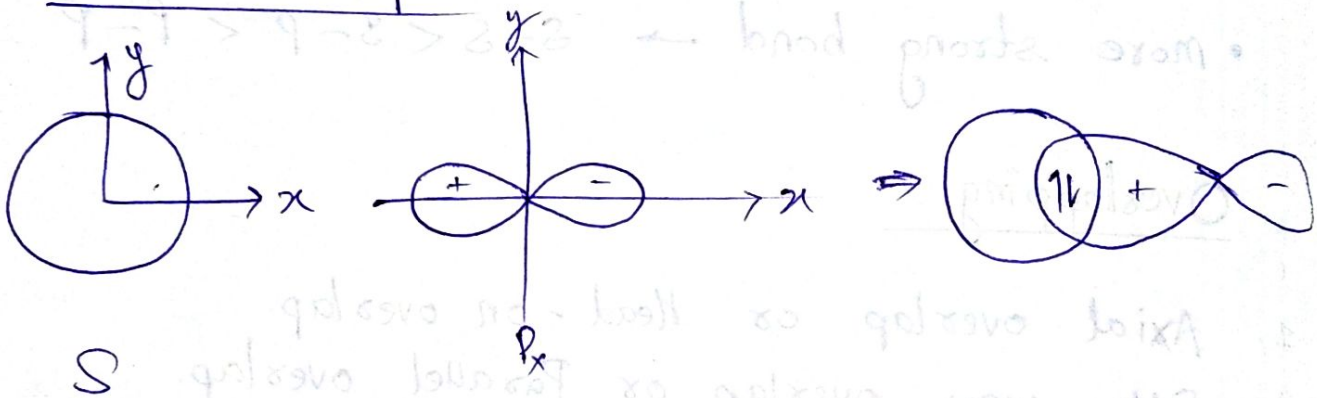
- (s-s overlap, s-p overlap, p-p overlap)

S – S Overlap



Eg. → H_2

S – P overlap



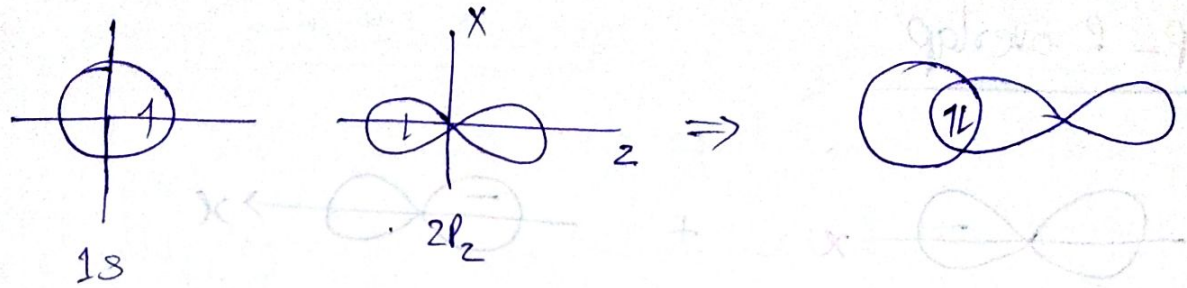
Eg → HF, HCl, HBr, H_2S, NH_3

HF

H	F
1	9
$1s^1$	$1s^2\ 2s^2\ 2p^5$
[↑]	[↑↓] [↑↓] [↑]
1s	$2P_x$ $2P_y$ $2P_z$

[↑] + [↑]

1s $2P_z$



g region

Eg. H_2S

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

↑↓	↑	↑

$H \rightarrow 1s^1$ [↑]

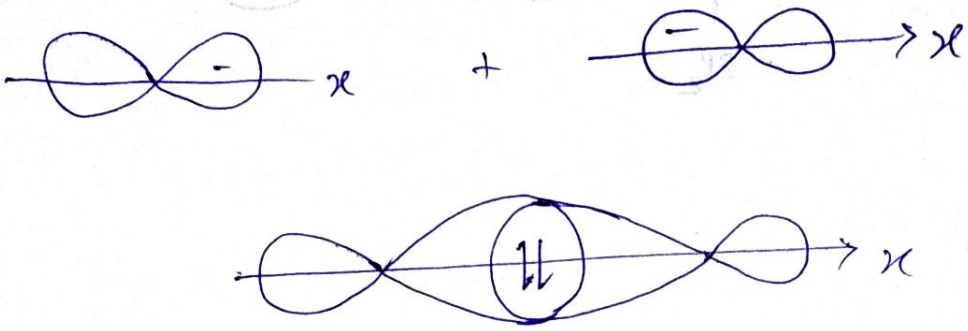
$H \rightarrow 1s^1$ [↑]

$H - \ddot{\underset{..}{S}} - H$

s-p overlap

s-p overlap

P – P overlap



$P_x - P_x \rightarrow x$ is inter nuclear axis (INA)
$P_y - P_y \rightarrow y$ is INA
$P_z - P_z \rightarrow z$ is INA

} axial overlap or Head on overlap

eg. $\rightarrow F_2, Cl_2, Br_2, I_2$

F_2

$_9F \rightarrow 1s^2 \quad 2s^2 \quad 2p^5$

$1s^2$	$2s^2$	$2p_x$	$2p_y$	$2p_z$
↑↓	↑↓	↑↓	↑↓	↑

$_9F \rightarrow$

		$2p_x$	$2p_y$	$2p_z$
↑↓	↑↓	↑↓	↑↓	↑

In $F_2 \rightarrow 2P - 2P$ overlap

similarly, In $Cl_2 \rightarrow 3p - 3p$

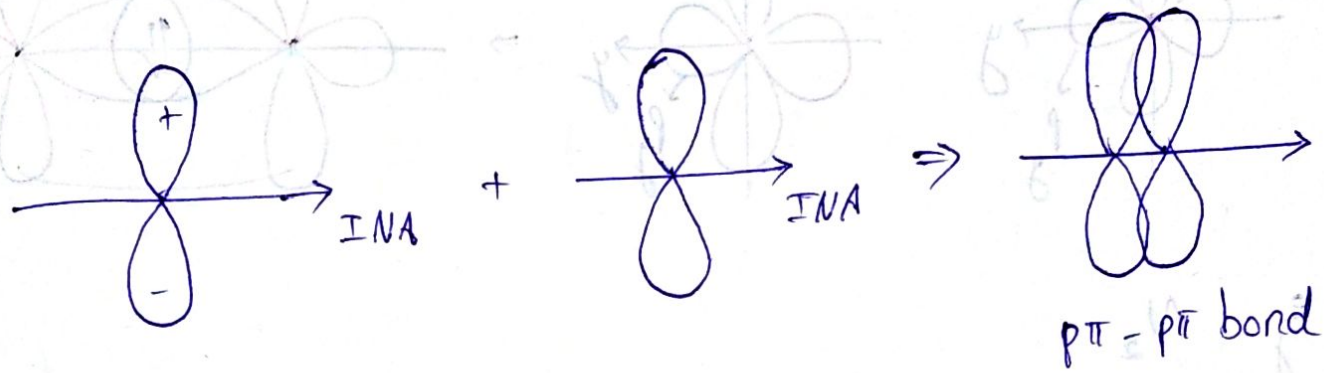
$Br_2 \rightarrow 4p - 4p$

$I_2 \rightarrow 5p - 5p$

Parallel Overlap

(p-p, p-d, d-d overlap)

- Overlapping of lobes above and below inter-nuclear axis is parallel or side-ways overlap.



Parallel overlap ⇒ p bond formation

- N_2, O_2

Eg: O_2 $\quad \ddot{O} \underset{\sigma}{\overset{\pi}{=}} \ddot{O}$

4 lone pair

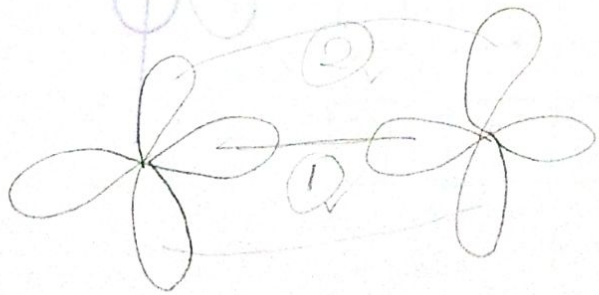
$O \rightarrow 1s^2\ 2s^2\ 2p^4$ $\quad$ [↑↓] [↑↓ | ↑ | ↑]

$O \rightarrow 1s^2\ 2s^2\ 2p^4$ $\quad$ [↑↓] [↑↓ | ↑↓ | ↑↓]

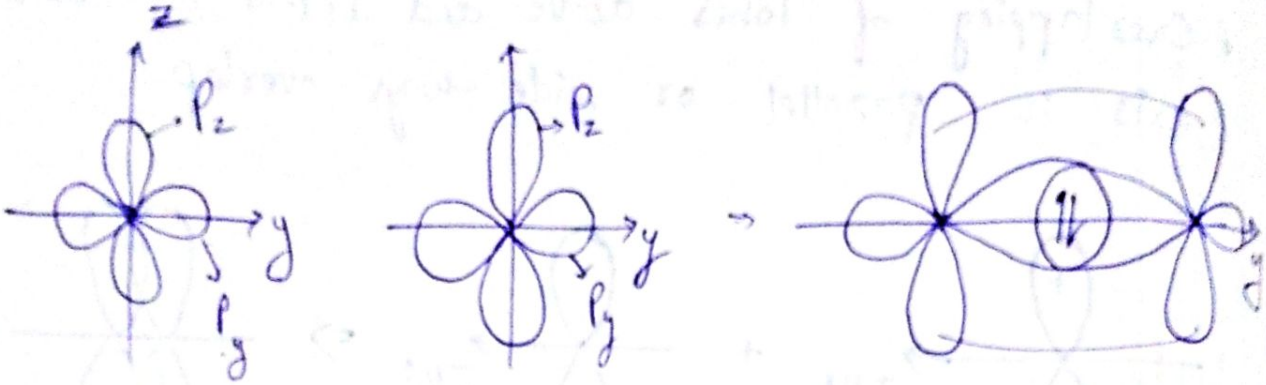
σ Bond $\quad$ π Bond

1st sigma bond is formed then, π bond is formed.

Reasons :-



In case of parallel overlap, first σ Bond is formed & then π bond is formed.



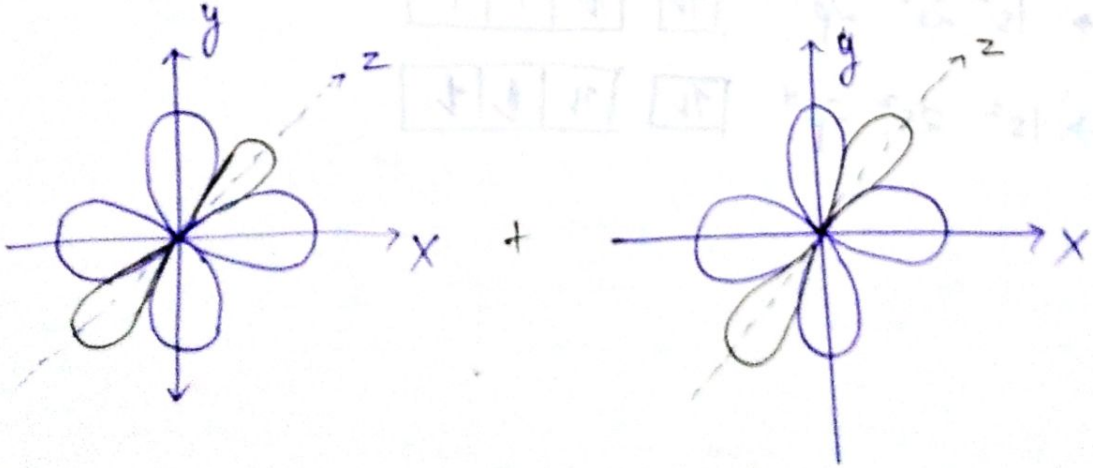
Eg N_2

$N \rightarrow 1s^2\ 2s^2\ 2p^3$

$N \rightarrow 1s^2\ 2s^2\ 2p^3$

1	1	1
$2P_x$	$2P_y$	$2P_z$

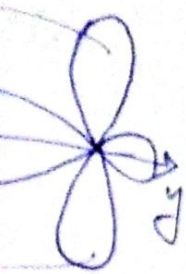
1	1	1
$2P_x$	$2P_y$	$2P_z$



$2P_x - 2P_x \rightarrow \sigma$ Bond (Axial overlap)

$2P_y - 2P_y \rightarrow p$ Bond (parallel overlap)

$2P_z - 2P_z \rightarrow p$ Bond (parallel overlap)

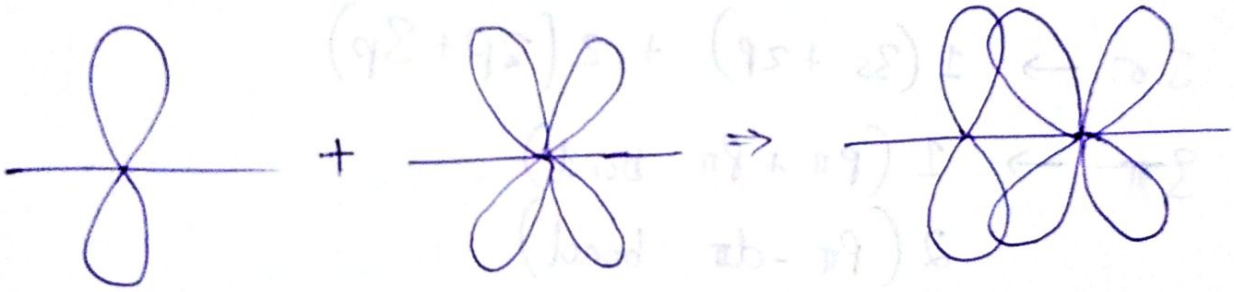


So, in $N_2 \rightarrow 1\sigma$ & $2p$ bond is formed

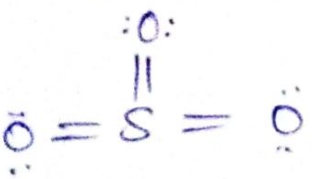
i) $1\ \sigma$ Bond

ii) $2\ P_\pi - P_\pi$ bond

P - d overlap

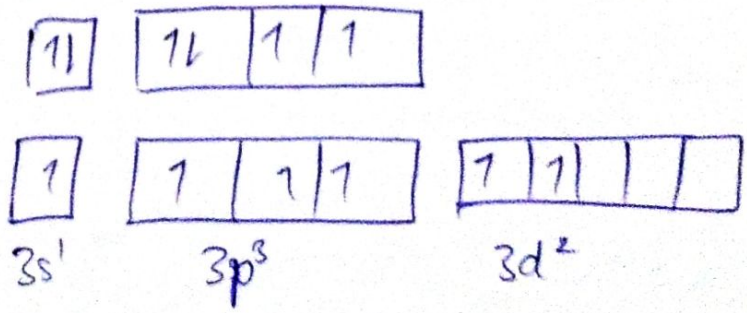


Eg. SO_3

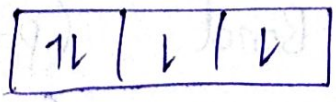


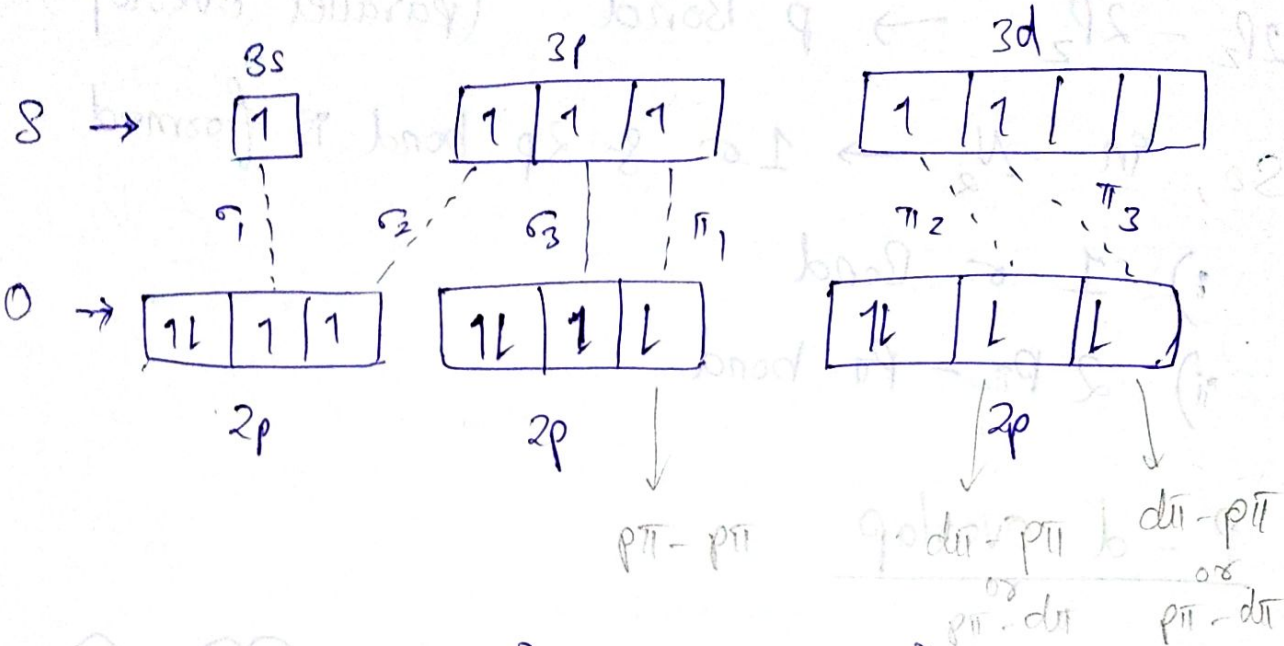
(3σ & 3π Bonds)

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



$_8O \rightarrow 1s^2 \; 2s^2 \; 2p^4$



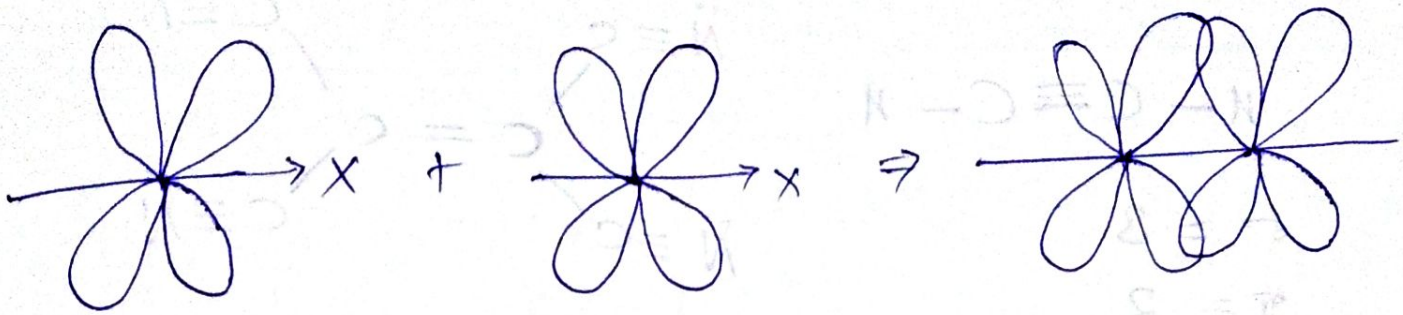


$3\sigma \rightarrow 1(3s + 2p) + 2(2p + 3p)$

$3\pi \rightarrow 1(p\pi + p\pi \text{ bond})$

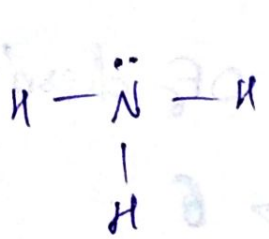
$2(p\pi - d\pi \text{ bond})$

D - D Overlap



Q Find out number of σ, π & lone-pair.

i) NH_3



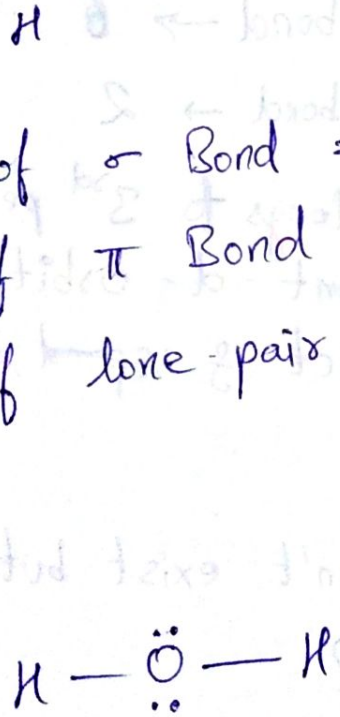
$\boxed{\text{No. of Bond pair} + \text{lone pair} = 4}$

∴ No. of σ Bond = 3

No. of π Bond = 0

No. of lone-pair = 1

ii) H_2O

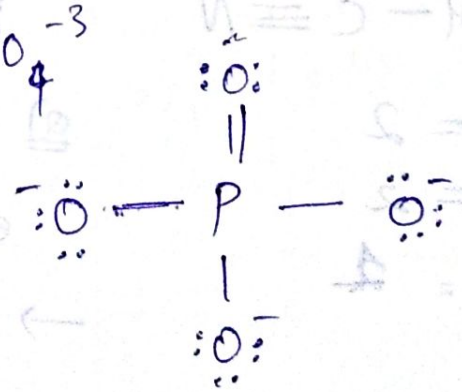


∴ σ = 2

π = 0

L.P = 2

iii) PO_4^{-3}



∴ σ = 4

π = 1

L.P = 11

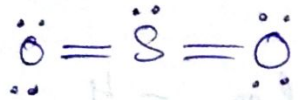
3) C_2H_2

$\sigma = 3$

$\pi = 2$

$L.P = 0$

4) SO_2



$\sigma = 2$

$\pi = 2$

$L.P = 5$

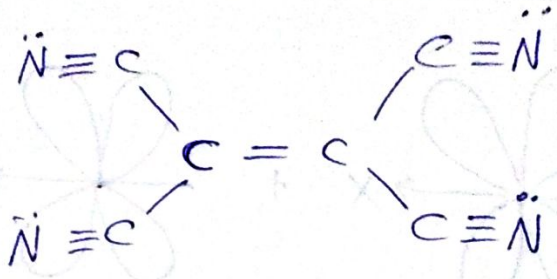
5) HCN

$\sigma = 2$

$\pi = 2$

$L.P = 1$

6) $C_2(CN)_4$



$L.P = 4$

$\sigma = 9$

$\pi = 9$

Q) SF_6 exists but OF_6 doesn't why?

→ S → max bond → 6

O → max bond → 2

also, S → belongs to 3rd period it has vacant d-orbital, so it can exict & ~~spand~~ expand the oclect.

Q) F_3^- doesn't exist but Cl_3^- / Br_3^- / I_3^- exist why?

→ Cl_3^- / Br_3^- / I_3^- are after 2nd period So they have vacant d-orbital hence they can expand the oclet but In F_3^- → it is 2nd period so it has p-orbital & cannot expand the oclet.

1st peri

2nd pe

3rd pe

4th p

5th p

Q) N

→ P

1st period → 1s
2nd period → 2s 2p
3rd period → 3s 3p 3d
4th period → 4s 4p 4d 4f
5th period → 5s 5p 5d 5f
⋮

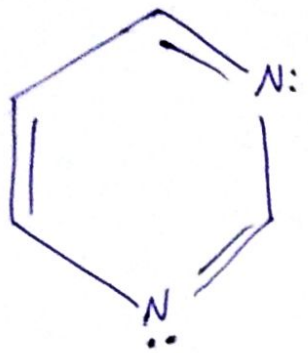
Q) NCl_5 doesn't exist but PCl_5 exist why?

→ P → 3rd period belong → so it has vacant d-orbital → so it can expand octet → and form 5 Bonds with Chlorine

N → 2nd period Belong → so it has No vacant d-orbital → So it cannot expand octet → & cannot form 5 Bonds with Chlorine

SKB → (Sir ki Baat)

Resonance in pyridinum



Here Resonance is <u>NOT</u> HAPPEN. But why?

Resonance happens in plane. whereas the lone pair is $\perp^{rd}$ to the plane here.

∴ Resonance will not happen in pyridinum.

HYBRIDIZATION

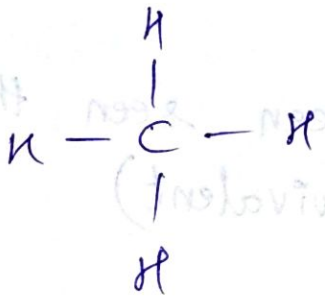
Lecture - 7

Q): Hybridization kya hai?

- Hybridization is an expansion of the V.B.T
- Intermixing of atomic orbital & formation of new hybrid orbital is called hybridization.

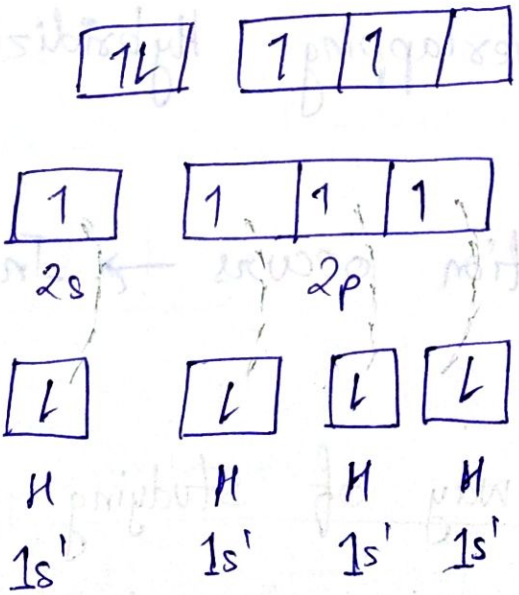
Why hybridization?

eg. CH_4



VBT in CH_4

$^{12}_{6}C \rightarrow 1s^2\ 2s^2\ 2p^2$



Theoretically,

We see that in CH_4 molecule, there is

1 → S-S Bond

3 → S-p Bond

A/c VBT 3 Bonds are of same type & one bond is different type.

Experimentally,

- Experimentally it has been seen that all 4 bonds are same (equivalent)

Hybridisation

- Before overlapping Hybridization occurs in molecule.

Hybridization occurs → Intermixing of orbitals occur.

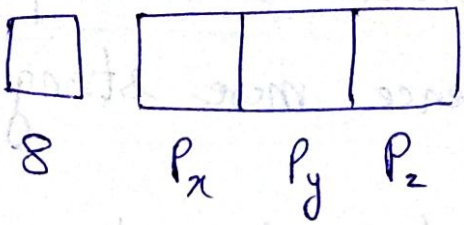
Bahubali way of studying

First intermixing of orbitals occurs

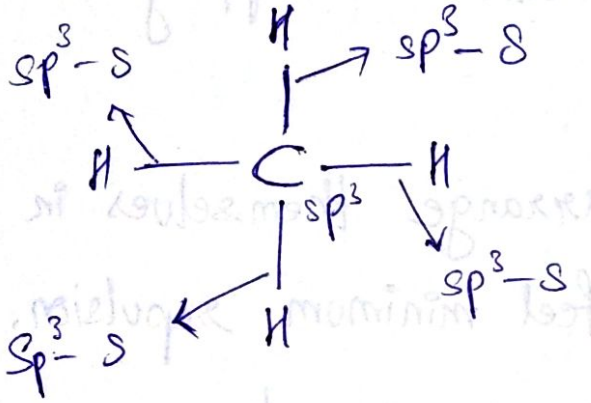
& then

overlapping occurs

$^{12}_{6}C \rightarrow 1s^2\ 2s^2\ 2p^6$



Hybridisation says that 1s orbital and 3p orbital combine together to form sp^3 hybrid orbital.



Ye hai Hybridization

- Therefore all bonds in CH_4 are equal since all 4 bonds are sp^3-s overlapping.

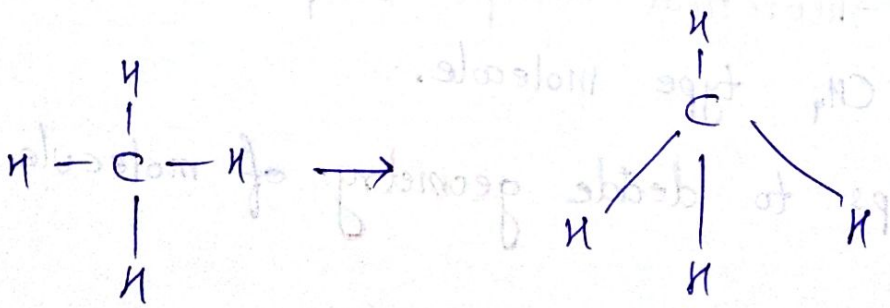
VIP Points in hybridization

- It is a theoretical concept to prove all bonds are same in CH_4 type molecule.
- It also helps to decide geometry of molecule

- Hybrid orbitals formed possess lower energy than pure orbital & thus provide better overlapping & hence more strong bond formed.
- Electrons do not participate in hybridization only orbitals take part in hybridization.
- Hybridized orbital show axial overlapping i.e. σ bonds are formed.
- The hybridized orbitals arrange themselves in such a way that they feel minimum repulsion. Hence they provide a proper geometry.

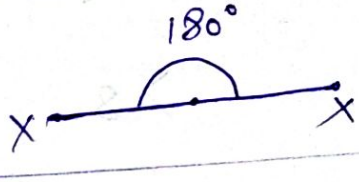
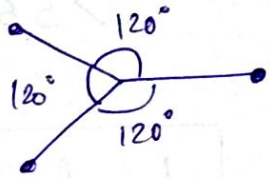
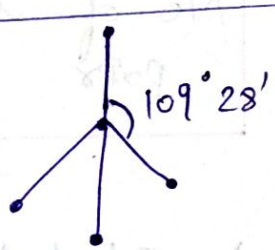
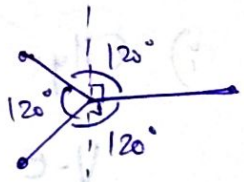
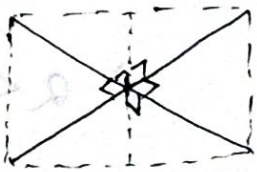
<u>After hybridization</u>

- All 4 bonds are sp^3-s hybridized
- So, all 4 hybridized orbital will adjust themselves for minimum repulsion.



(Tetrahedral)

Geometry on the Basis of Hybridization

Total number of Hybridized orbitals	Eg.	Hybridization	Geometry	
2	$BeCl_2$, $HgCl_2$, BeF_2	sp	Linear	
3	BF_3, NO_3^-, CO_3^{2-}	sp^2	Trigonal planar	
4	CH_4, NH_4^+, CCl_4, SiH_4, XeO_4	sp^3	Tetra-hedral	
5	PCl_5, AsF_5, $SbCl_5$, XeO_3F_2	sp^3d	Trigonal bipyramidal	 $3 \rightarrow 120°$ $2 \rightarrow 90°$
6	SF_6, $SnCl_6^{2-}$	sp^3d^2	Octahedral	 All 6 angle 90°
7	IF_7, $[ZrF_7]^{3-}$	sp^3d^3	Pentagonal bi-pyramidal	

TRICK

① Calculate sum of valence electron in molecule :

$> 2 \quad \leq 8 \rightarrow \div 2$

$> 8 \quad \leq 56 \rightarrow \div 8$

$> 56 \rightarrow \div 18$

No. of lone pair = Hybridized orbitals − Surrounding atom

Q) Calculate hybridization, Geometry, Lone pair

i) CH_4

V.e = valence e^- of C + 4 × valence of H

= 4 + 4 × 1

ve = 8

$\because 2 < V.e \leq 8 \Rightarrow \therefore \frac{8}{2} = 4 \Rightarrow \boxed{sp^3}$

$sp^3 \Rightarrow$ $\boxed{\text{Tetrahedral}}$

L.P = Hybrizided orbitals − S. atoms = 4 − 4 = $\boxed{0}$

ii) ClO_4^-

17^{Cl} → 2, 8, $\boxed{7}$ $\quad$ $_8O$ → 2, $\boxed{6}$

$v.e = 7 + 4 \times 6 + 1$

$= 7 + 24 + 1$

$= 32$

$\therefore 8 < v.e \leq 56 \Rightarrow \div 8 \Rightarrow \frac{32}{8} = 4 \Rightarrow \boxed{sp^3}$

$sp^3 \Rightarrow$ $\boxed{\text{Tetrahedral.}}$

No. of $L.P = 4 - 4 = \boxed{0}$

iii) SO_2

16^{S} → 2, 8, $\boxed{6}$ $\quad$ $_8O$ → 2, $\boxed{6}$

$v.e^- = 6 + 2 \times 6$

$= 18$

$\because 8 < v.e \leq 56 \Rightarrow \div 8 \Rightarrow \frac{18}{8} \Rightarrow$ 8)18 (2, 16, 2)2 (1, 2, ×) → $\boxed{3} \Rightarrow \boxed{sp^2}$

$sp^2 \rightarrow$ $\boxed{\text{Trigonal planar}}$

No. of $L.P = 3 - 2 = \boxed{1}$

iv) CO_2

$^{12}_{6}C$ $^{16}_{8}O$

$\Rightarrow$ V.e $= 4 + 6\times2$
$= 4 + 12$
$= 16$

$\because$ $8 < ve \leq 56 \Rightarrow \div 8$

$\therefore$

$8\overline{)16}(2$
$\underline{16}$
$(\times)$

$2 \Rightarrow \boxed{sp} \Rightarrow \boxed{\text{linear}}$

L.P. $= 2 - 2 = \boxed{0}$

v) NH_4^+

$^{14}_{7}N$ $^{1}_{1}H$

$\Rightarrow$ V.e $= 5 + 1\times4 - 1$
$= 5 + 4 - 1$
ve $= 8$

$\because$ $2 < V.e \leq 8 \Rightarrow \div 2$

$\therefore$ $2\overline{)8}(4$
$\underline{8}$
(0)

$\therefore 4 \Rightarrow sp^3 \Rightarrow$ Tetrahedral

L.P $= 4 - 4 = 0$

vi) ClF_3

$^{35.5}_{17}Cl$ $^{19}_{9}F$

$\Rightarrow$ V.e $= 7 + 3\times7$
$= 7 + 21 = 28$

Since $8 < Ve \leq 56 \Rightarrow \div 8$

$\therefore$ $8\overline{)28}(3$
$\underline{24}$
$2\overline{)4}(2$
$\underline{4}$
(0)

$2 < 4 \leq 8$
$\div 2$

$5 \Rightarrow \boxed{sp^3d}$

$sp^3d \Rightarrow \boxed{\text{Trigonal bipyramidal}}$

L.P $= 5 - 3 = \boxed{2}$

vii) XeF_6

$\Rightarrow$ V.e $= 8 + 7\times6$
$= 8 + 42 = 50$

$\because$ $8 < Ve \leq 56 \Rightarrow \div 8$

$8\overline{)50}(6$
$\underline{48}$
$2\overline{)2}(1$
$\underline{2}$
$(\times)$

$7 \Rightarrow \boxed{sp^3d^3}$

$\therefore sp^3d^3 \Rightarrow \boxed{\text{Pentagonal bipyramidal}}$

L.P $= 7 - 6 = \boxed{1}$

viii) I_3

$\Rightarrow$

$\therefore$ 8

$\therefore$ sp

L.

Hybr

N

viii) I_3^-

$\Rightarrow$ $V.e = 3 \times 7 + 1$

$= 22$

$\therefore$ $8 < V.e \leq 56 \Rightarrow \div 8$

$\therefore$

$$\begin{array}{r} 8\overline{)22}\,(2 \\ 16 \\ \hline 2\overline{)6}\,(3 \\ 6 \\ \hline 0 \end{array} \Big\rangle\ 5 \Rightarrow sp^3d$$

$\therefore$ $sp^3d \Rightarrow$ Trigonal bipyramidal

$L.P \Rightarrow$ By diagram $\rightarrow$

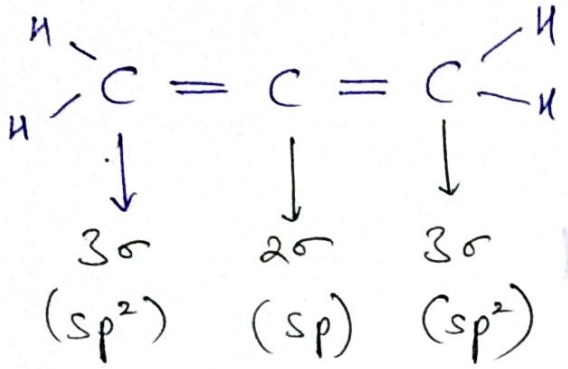
I

Hybridisation for oxy-acid / molecule having more than one central atom.

No. of σ Bond	Hybridisation
2	sp (linear)
3	sp^2 (Trigonal planar)
4	sp^3 (Tetrahedral)
5	sp^3d (T.B.P
6	sp^3d^2 (octahedral)
7	sp^3d^3 (P.B.P)

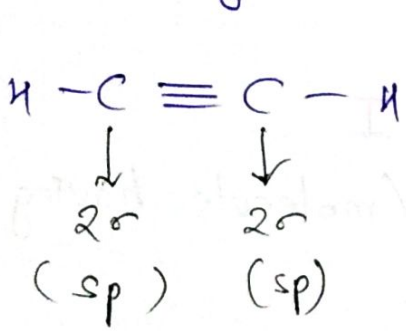
Eg. Calculate hybridization of all carbon atom in allene.

⇒ • Allene → C_3H_4



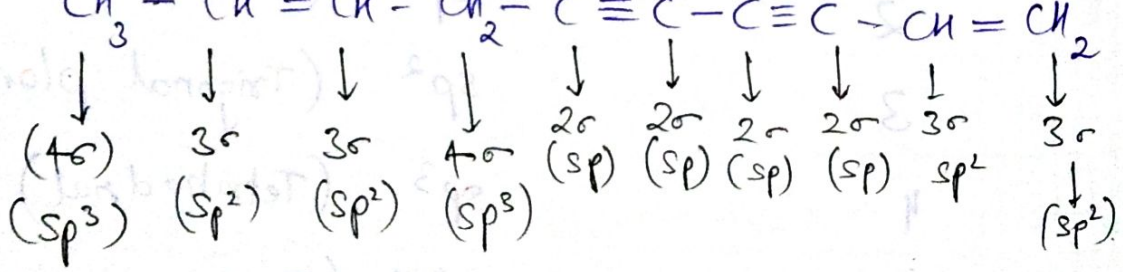
Eg. C_2H_2 (ethyne)

→



Eg. Cal. hybridisation for each carbon atom.

→



V.S.

Hybridiz

i) CH_4

v.e = 4 +

= 8

2) 8 (

8

(×)

sp^3 → Te

L.P = 4 -

Accord

CH_4 →

NH_3 —

H_2O —

Exper

CH_4 →

V.S.E.P.R Theory (Valence shell electron pair Repulsion Theory)

Hybridization & L.P Nikalo.

i) CH_4

v.e = 4 + 1×4
= 8

2) 8 (4 → sp^3
8
(×)

sp^3 → Tetrahedral

L.P = 4 − 4 = 0

ii) NH_3

v.e = 5 + 3×1
= 8

2) 8 (4 → sp^3
8
(×)

sp^3 → Tetrahedral

L.P = 4 − 3 = 1

iii) H_2O

v.e = 2×1 + 6 = 2+6
= 8

2) 8 (4 → sp^3
8
(×)

sp^3 → Tetrahedral

L.P = 4 − 2 = 2

According to VBT

CH_4 → sp^3 → Tetrahedral → 109°28'

NH_3 → sp^3 → Tetrahedral → 109°28'

H_2O → sp^3 → Tetrahedral → 109°28'

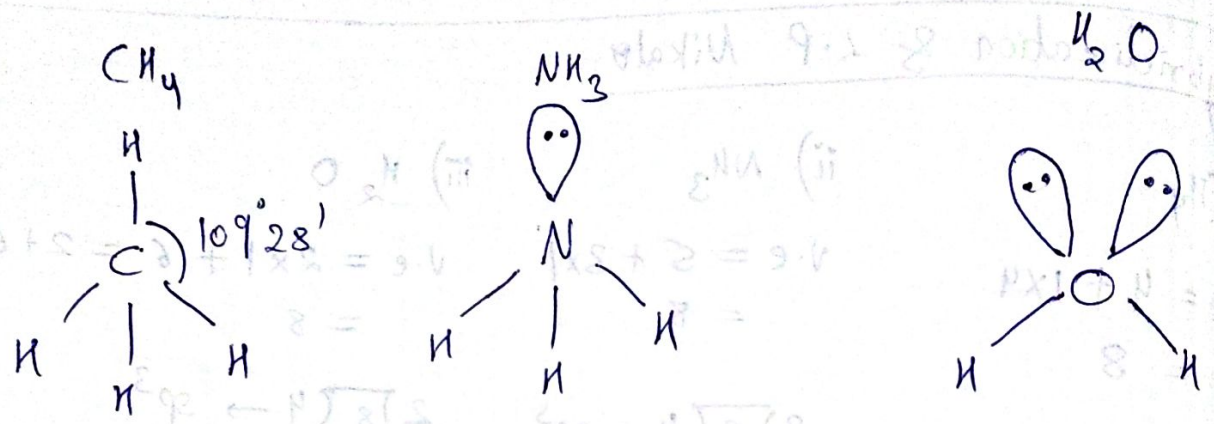
Experimentally

CH_4 → sp^3 ⇒ 109° 28' → L.P = 0

NH_3 → sp^3 ⇒ 107° → L.P = 1

H_2O → sp^3 ⇒ 104° → L.P = 2

Why bond angle is reduced?



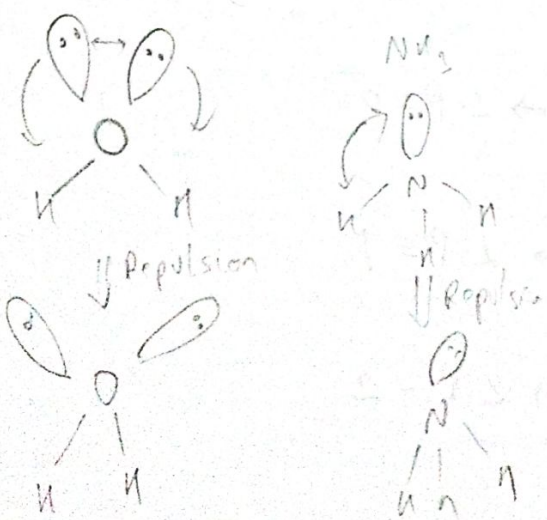
VSERR Theory

VSEPR says that there are two types of electron pair in a molecule.

1) Bond pair
2) Lone pair

repulsion order is –

$$\boxed{LP - LP > LP - BP > BP - BP}$$



Why lone pair has maximum repulsion ?

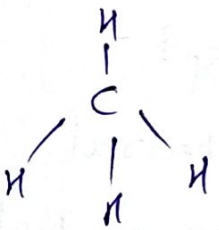
- Lone pair has maximum repulsion and bond pair has less repulsion.

Because

In Bond pair electrons are divided between two atoms, & in lone pair electrons are at same place & hence repulsion is more.

$$\boxed{Lp - Lp > Lp - Bp > Bp - Bp}$$

CH_4



less repulsion due to bond pair - bond pair repulsion

$\approx 109°$

NH_3

one lone pair so repulsion is more than first one

$\approx 107°$

H_2O

max repulsion due to two lone pair

$\approx 104°$

Geometry v/s shape

- Geometry of a molecule depends on (lone pair & bond pair both)
- Shape of a molecule depends on only bond pair

sp^2

Number of Bonding

3

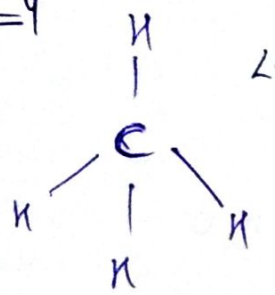
2

sp^3

4

3

CH_4



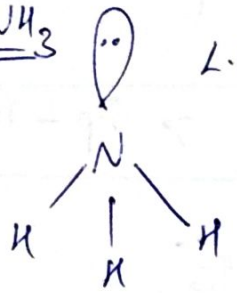
L.P = 0

sp^3

Geometry :- Tetrahedral

Shape :- Tetrahedral

NH_3



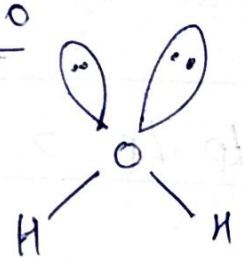
L.P = 1

sp^3

Tetrahedral

Pyramidal

H_2O



L.P = 2

sp^3

Tetrahedral

Angular or Bent or Opposite V shape

Shape of Molecule in VSEPR

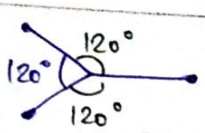
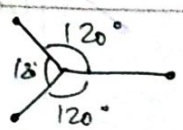
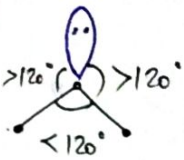
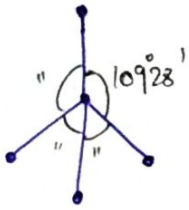
e pair

ed pair

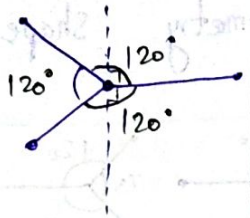
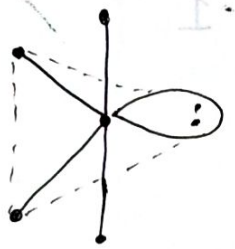
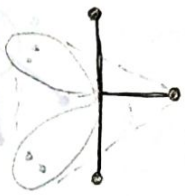
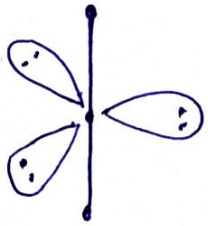
L.P = 2

ent or

shape

Number of Bonding pair	No. of L.P on C.M.A	Geometry	Shape	eg
sp^2				
3	0	 Trigonal planar	 Trigonal Planar	BF_3, BCl_3, C_2H_4, CO_3^{2-}, NO_3^-
2	1	"	 Angular/Bent /V shaped	SO_2, O_3, $SnCl_2$
sp^3				
4	0	 Tetrahedral	Tetrahedral	CH_4, SiF_4, NH_4^+, ClO_4^-
3	1	"	Trigonal pyramidal	NH_3, PCl_3
2	2	"	Angular/ Bent	H_2O, OF_2, H_2S, SCl_2

sp^3d

B.P	L.P	Geometry	Shape	Eg.
5	0	 $3 \rightarrow 120°$ $2 \rightarrow 90°$ Trigonal bipyramidal	→ " T.B.P	PF_5, PCl_5
4	1	" TBP	 See-Saw	SF_4
3	2	" TBP	 T-shaped	ClF_3, BrF_3, ICl_3
2	3	" TBP	 (Linear)	XeF_2

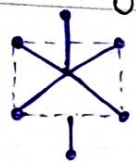
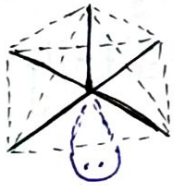
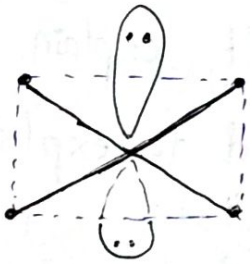
sp^3d^2

B.P
6
5
4

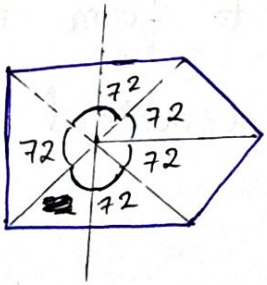
sp^3d^3

7

sp^3d^2

B.P	L.P	Geometry	Shape	Eg
6	0	 All 12 Bond Angle = 90° Octahedral	→ "	SF_6, $[CrF_6]^{3-}$, $[Co(NH_3)_6]^{3+}$
5	1	"	 Square - pyramidal	BrF_5, IF_5
4	2	"	 (Square-planar)	XeF_4

sp^3d^3

7	0		"	IF_7

MOT (Molecular Orbital Theory)

Why MOT ?

- VBT & VSEPR can't explain some properties of molecules

Such as

- A/c VBT O_2 is dia-magnetic
  But experimentally O_2 is paramagnetic

VIP points

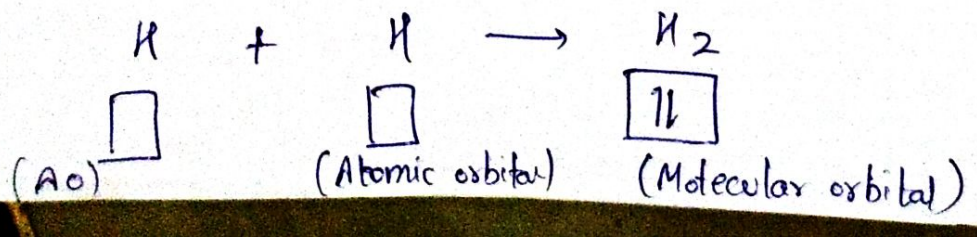
1. Paramagnetic nature of O_2 is explained by MOT
2. VBT can't explain about fractional bond order.
3. VBT failed to explain stability of molecule.

   e.g O_2, O_2^+, O_2^-

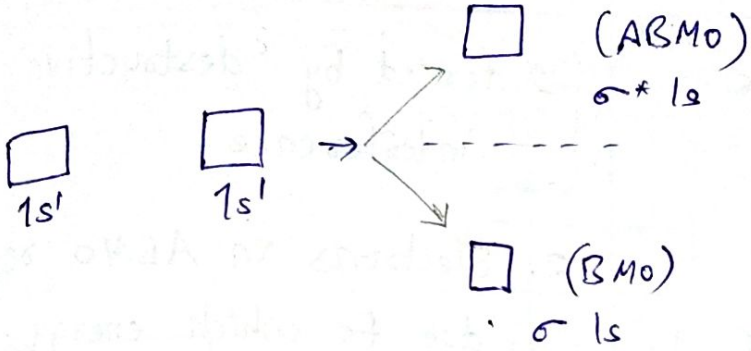
MOT explains above things easily.

A/c MOT

- Atomic orbitals Combines to form molecular orbital
- No. of atomic orbitals Combined = no. of molecular orbitals formed.



$H + H \rightarrow H_2$



Atomic orbital → s, p, d, f

Molecular orbital → σ, π, δ

There are two types of Molecular orbitals:

i) Bonding Molecular orbital (BMO)

ii) Anti-Bonding Molecular orbital (ABMO)

BMO → less energy, more stable

ABMO → high energy, less stable

BMO & ABMO

BMO	ABMO
1. formed by constructive interference	1. formed by destructive interference
2. Electrons in BMO attract each other due to which energy↓ & hence stability↑	2. Electrons in ABMO repel, due to which energy increase & stability↓
3. Lower energy than atomic orbital hence most stable	3. Higher energy than atomic orbital hence least stable

Combing Atomic orbitals

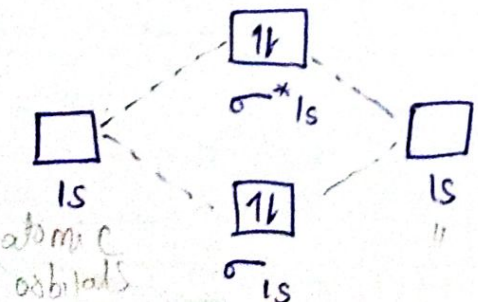
Eg. N_2

$\Rightarrow N + N \longrightarrow N_2$

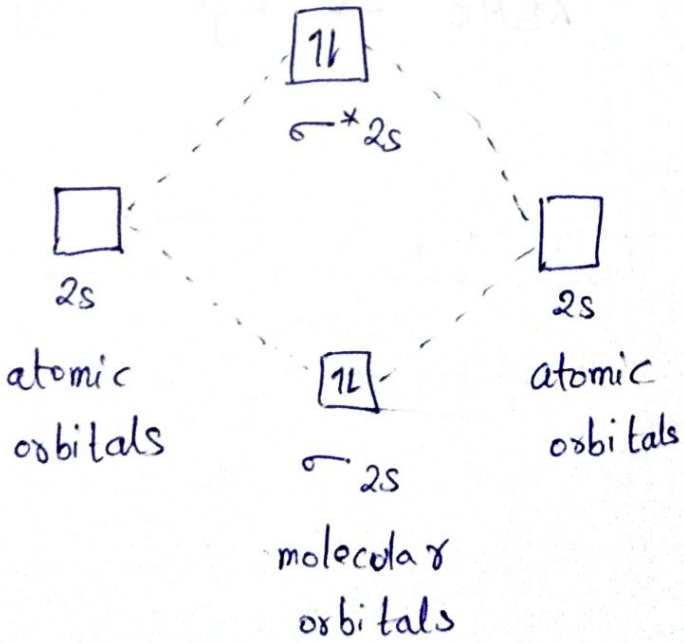
EC: $^{14}_{7}N \rightarrow 1s^2\ 2s^2\ 2p^3$

$N \rightarrow 1s^2\ 2s^2\ 2p^3$

Firstly, 1s & 1s will combine & form BMO & ABMO.

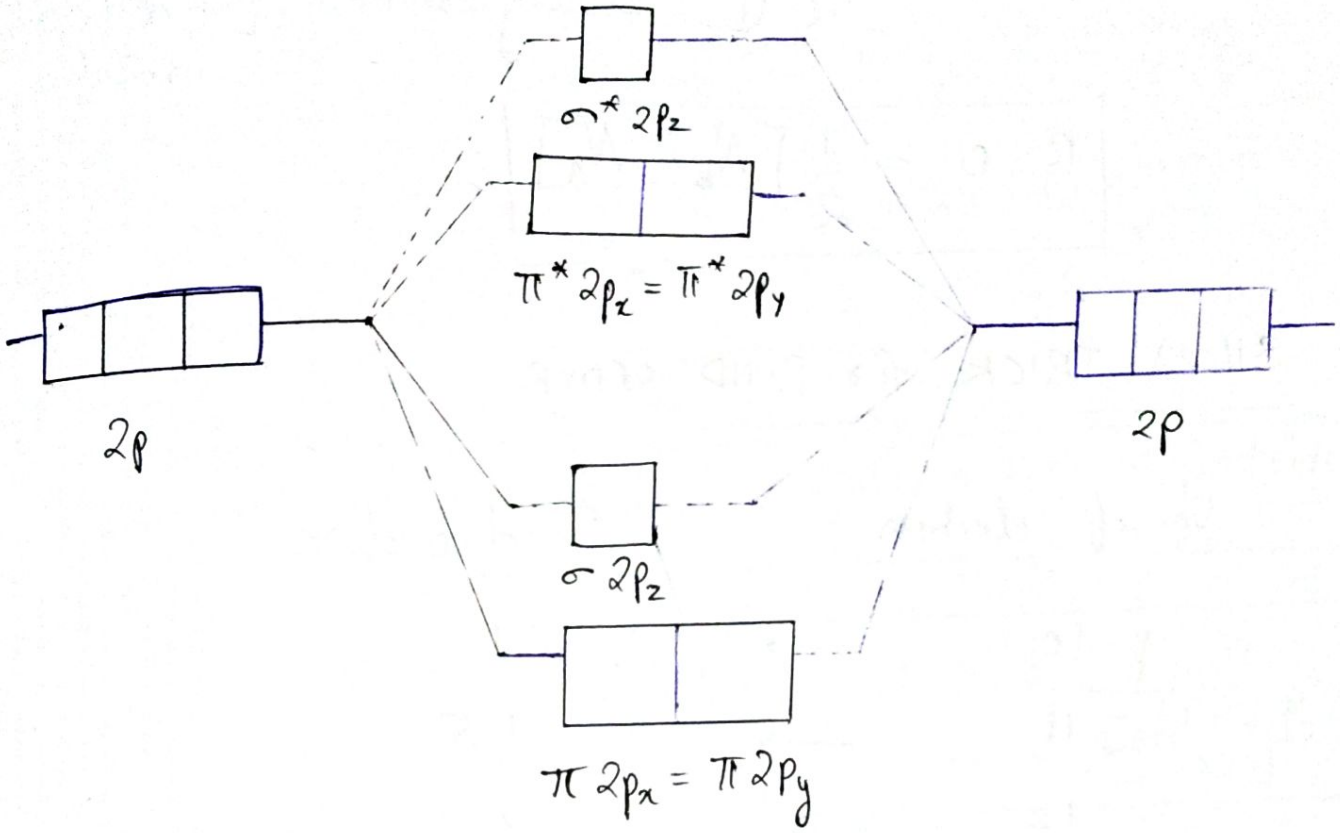


2s & 2s will combine

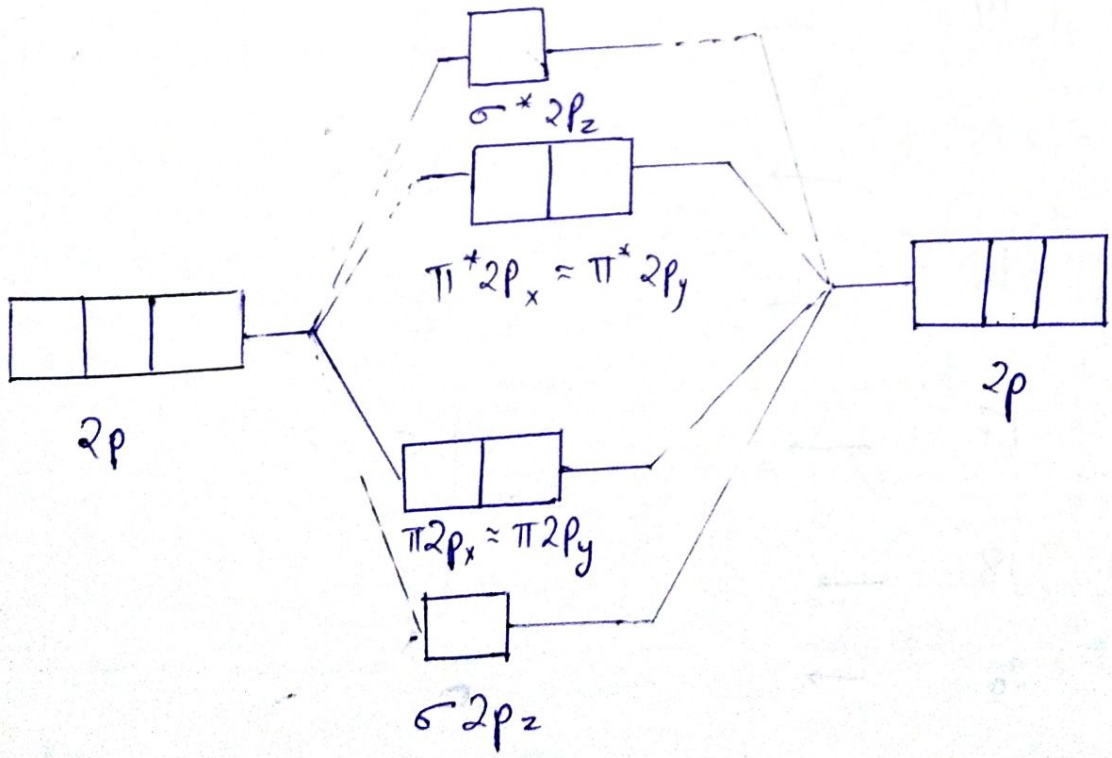


2p & 2p will combine. Two cases are these

i) Total electrons <= 14



ii) Total electrons >14



Bond order

$$\text{Bond order} = \frac{1}{2}\left\{\left[\begin{array}{c}\text{No. of electrons}\\ \text{in Bonding orbitals}\end{array}\right] - \left[\begin{array}{c}\text{No. of electrons}\\ \text{in Anti-Bonding}\\ \text{orbitals}\end{array}\right]\right\}$$

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

SHORT TRICK for BOND ORDER

No. of electron		Bond order
10	→	1
11	→	1.5
12	→	2
13	→	2.5
14	→	3
15	→	2.5
16	→	2
17	→	1.5
18	→	1
20	→	0

Eg. Cal. B.O

i) O_2

→ No. of e⁻

∴ 16e⁻ —

ii) CO^+

→ No. of e

∴ 13e⁻ —

i.e C≡

But it

B.O of

iii) O_2^+

→ No. of

∴ 15e⁻

iv) B_2

→ No.

10e⁻

Eg. Cal. B.O

lectrons Bonding

i) O_2

→ No. of e^- = $8 \times 2 = 16e^-$

∴ $16e^- \longrightarrow$ B.O = 2

ii) CO^+

→ No. of e^- = $6 + 8 - 1$
= $13e^-$

∴ $13e^- \longrightarrow$ B.O = 2.5

i.e $C \equiv O^+$

But it is exception

B.O of $CO^+ > 3$

iii) O_2^+

→ No. of e^- = $2 \times 8 - 1$
= $16 - 1$
= $15e^-$

∴ $15e^- \longrightarrow$ B.O = 2.5

iv) B_2

→ No. of e^- = $2 \times 5 = 10e^-$

$10e^- \longrightarrow$ B.O = 1

5) N_2

→ No. of e^- = $2 \times 7 = 14e^-$

∴ $14e^- \longrightarrow$ B.O = 3

6) O_2^-

→ No. of e^- = $2 \times 8 + 1 = 17e^-$

∴ $17e^- \longrightarrow$ 1.5 = B.O

Note:

Bond dissociation energy ∝ Bond order

B.O ∝ B.E ∝ stability ∝ $\frac{1}{\text{Bond length}}$

B.E = Bond Energy

- B.O = 1 → i.e (Single Bond)
- B.O = 2 → i.e (double Bond)
- B.O = 3 → i.e (triple Bond)

B.O = fraction → 1.5 → [A ⋯ A]
→ 2.5 → [A ⋯ A]
→ 3 → [A ≡ A]

Stability

If ~~Bonding~~ Bond order of two compounds are same then check for number of electrons in BMO & ABMO.

More no. of electrons in ABMO $\Rightarrow$ Less stable

Diamagnetic & Para-magnetic

Sum of electrons:

Even = Dia-magnetic

odd = paramagnetic

Exception : B_2 & O_2
$B_2 \downarrow$ (para) $O_2 \downarrow$ (para)

Q) Calculate bond-order & tell the magnetic behaviour.

i) N_2

$\rightarrow$ No. of $e^- = 2 \times 7 = 14e^-$

$14e^- \rightarrow B.O = 3$
(even)

$\therefore N_2 \rightarrow$ dia-magnetic

ii) B_2

$\rightarrow$ No. of $e^- = 2 \times 5 = 10e^-$

$10e^- \rightarrow B.O = 1$
(even)

$\therefore B_2 \rightarrow$ dia-magnetic ✓

$B_2 \rightarrow$ para-magnetic (exception)

iii) C_2

$\rightarrow$ No. of

$12e^-$
(even)

$\therefore C_2$ —

iv) N_2^+

$\rightarrow$ No. o

$13e^-$
(odd)

$\therefore N_2^+$

(5) N_2^-

$\Rightarrow$ No.

$15e$
(odd

$\therefore N_2^-$

(6) O_2

$\Rightarrow$ No.

$16e^-$
(even

O_2

iii) C_2

→ No. of e^- = $2 \times 6 = 12e^-$

$12e^-$ (even) ⟶ $2 = B.O$

∴ C_2 → Dia-magnetic

iv) N_2^+

→ No. of e^- = $2 \times 7 - 1$
= $13e^-$

$13e^-$ (odd) ⟶ $2.5 = B.O$

∴ N_2^+ → paramagnetic

⑤ N_2^-

⇒ No. of e^- = $2 \times 7 + 1$
= $15e^-$

$15e^-$ (odd) ⟶ $2.5 = B.O$

∴ N_2^- → Paramagnetic

⑥ O_2 [Exception]

⇒ No. of e^- = $2 \times 8 = 16e^-$

$16e^-$ (even) ⇒ $B.O = 2$

O_2 ⟶ paramagnetic

⑦ O_2^+

⇒ No. of e^- ⇒ $2 \times 8 - 1 = 16 - 1 = 15e^-$

∴ $15e^-$ (odd) ⟶ $2.5 = B.O$

∴ O_2^+ → paramagnetic

⑧ O_2^{+2}

→ No. of e^- = $2 \times 8 - 2 = 16 - 2 = 14e^-$

∴ $14e^-$ (even) ⟶ $3 = B.O$

∴ O_2^{+2} → diamagnetic

⑨ O_2^-

→ No. of e^- ⇒ $2 \times 8 + 1 = 17e^-$

$17e^-$ (odd) ⟶ $B.O = 1.5$

∴ O_2^- → ~~dia~~ paramagnetic

⑩ O_2^{-2}

→ No. of e^- ⇒ $2 \times 8 + 2 = 18e^-$

$18e^-$ (even) ⟶ $B.O = 1$

∴ O_2^{-2} → diamagnetic

Eg. Write stability order N_2, N_2^+, N_2^-

→ $\boxed{B.O \propto \text{stability}}$

$N_2 \rightarrow B.O = 3$

$N_2^+ \rightarrow B.O = 2.5$

$N_2^- \rightarrow B.O = 2.5$

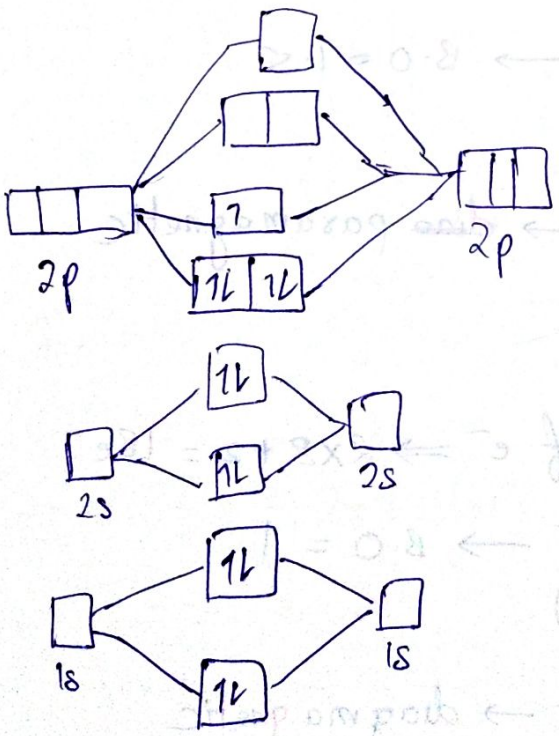
Now, N_2^+ & N_2^- have same Bond order.

$N_2^+ \rightarrow$ no. of e^-s $= 2 \times 7 - 1 = 13e^-$

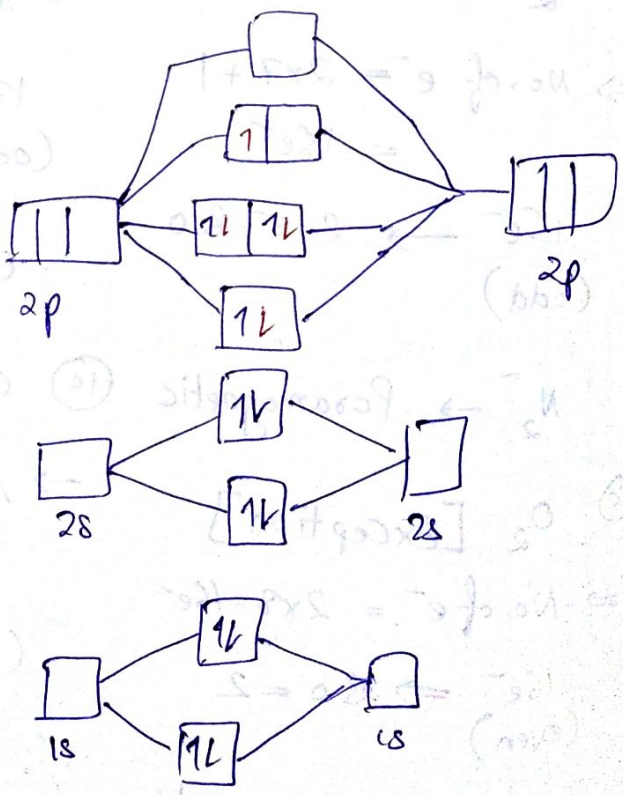
$N_2^- \rightarrow$ no. of e^-s $= 2 \times 7 + 1 = 15e^-$

$N \rightarrow 1s^2 2s^2 2p^3$

N_2^+ e^-s ≤ 14



N_2^- e^-s > 14



More the no. of e^-s in ABMO $\Rightarrow$ Less stable

$\therefore N_2 > N_2^+ > N_2^-$

Eg. write stability order & magnetic behavior:

O_2, O_2^+, O_2^{+2}, O_2^-, O_2^{-2}

→ Bond order of $O_2 = 2$ → (paramagnetic)

Bond " of $O_2^+ = 2.5$

Bond order of $O_2^{+2} = 3$

Bond order of $O_2^- = 1.5$

Bond " of $O_2^{-2} = 1$

$\boxed{\text{Bond order} \propto \text{stability}}$

$\therefore O_2^{+2} > O_2^+ > O_2 > O_2^- > O_2^{-2}$

Eg. Arrange following in O–O Bond-length.

KO_2 → O_2^-

Na_2O_2 → O_2^{-2}

O_2 → O_2

$O_2^{+2}[PtF_6]^{-2}$ → O_2^{+2}

$\boxed{\text{Bond order} \propto \frac{1}{\text{Bond length}}}$

	O_2^-	O_2^{-2}	O_2	O_2^{+2}
$e^- =$	$17e^-$	$18e^-$	$16e^-$	$14e^-$
B.O =	1.5	1	2	3

∴ Bond ~~length~~ order : $O_2^{+2} > O_2 > O_2^- > O_2^{-2}$

∴ Bond length order : $O_2^{-2} > O_2^- > O_2 > O_2^{+2}$

Eg. Which is more stable.

He_2^- , He_2^+

→ $\downarrow$ $\downarrow$

e^-s ~~3×2~~ $2\times2-1$

$2\times2+1$

$\downarrow$ $\downarrow$

e^-s $5e$ $3e$

He_2^-

B A B A

2 2 1

$$B.O = \frac{B-A}{2} = \frac{3-2}{2}$$

$$\boxed{B.O = \frac{1}{2}}$$

He_2^+

B A B A

2 1

$$B.O = \frac{B-A}{2} = \frac{2-1}{2} = \frac{1}{2}$$

$$\boxed{B.O = \frac{1}{2}}$$

More no. of e^-s in ABMO $\Rightarrow$ less stable

$\therefore He_2^+ > He_2^-$ } stability order

HOMO & LUMO

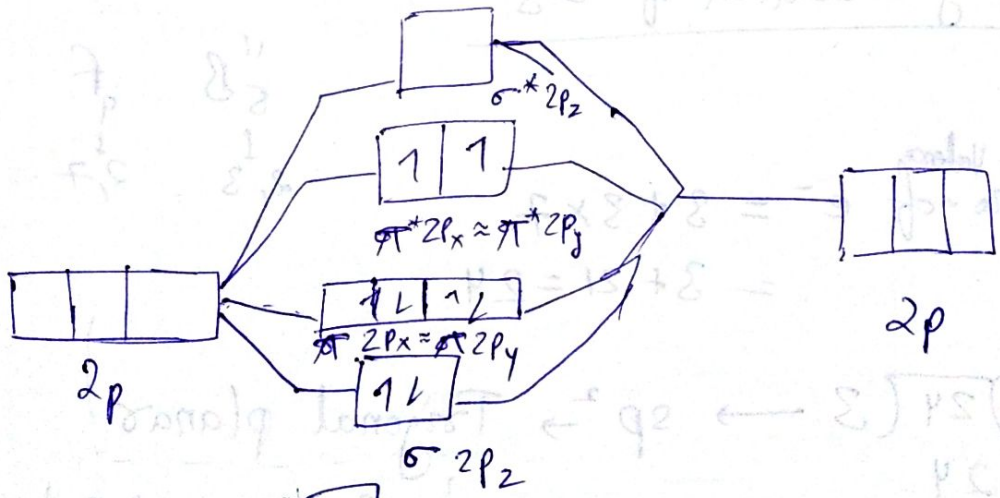
- HOMO $\rightarrow$ Highest occupied molecular orbital.
- LUMO $\rightarrow$ Lowest unoccupied molecular orbital

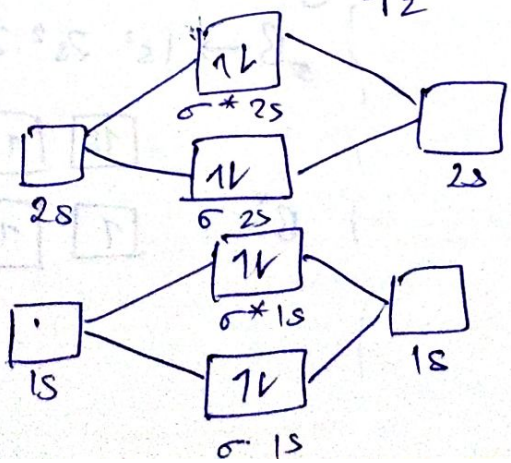
Eg. $O_2 \rightarrow$ Total no. of $e^- = 16e^-$

E. Conf $O = 1s^2\ 2s^2\ 2p^4$

$O = 1s^2\ 2s^2\ 2p^4$

Total electrons $> 14e^-$





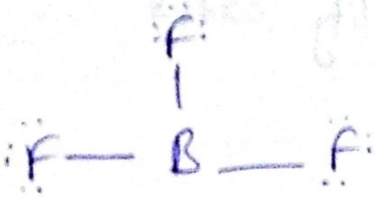
$\therefore$ HOMO $= \pi^* 2p_x \approx \pi^* 2p_y$

LUMO $= \sigma^* 2p_z$

BACK-BONDING

Draw lewis structure of BF_3

→



→ 1.50A° (assumed)
↳ Bond length

BF_3 is hypovalent

Origin of Back Bond

- Theoretically single Bond B-F bond length is ≃ 1.50A°
- Experimentally it has been found that B-F bond length in BF_3 is ≃ 1.30 A°

Calculate Hybridization of BF_3

- BF_3

$_5B$ → 2, 3 $\quad$ $_9F$ → 2, 7

⇒ Total Valence no of $e^- = 3 + 3 \times 7$
$= 3 + 21 = 24$

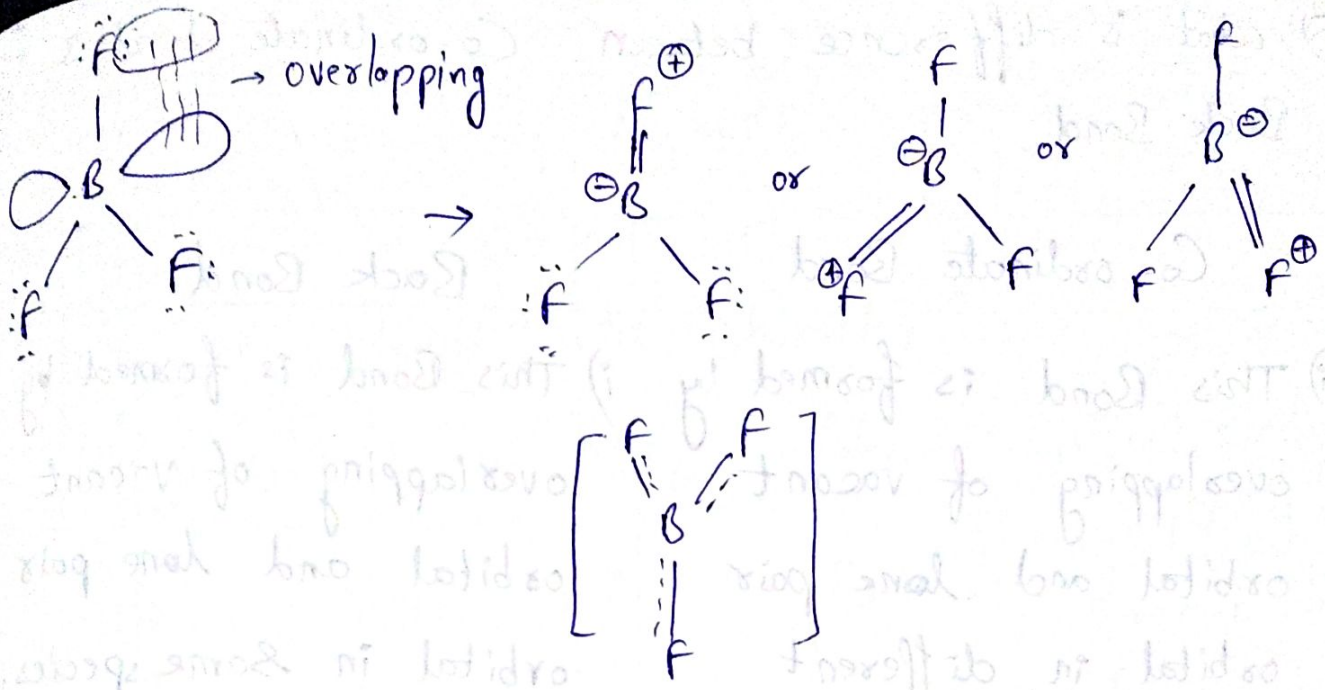
∴ $8\overline{)24}\,(3 \longrightarrow sp^2 \rightarrow$ Trigonal planar
24
(×)

$_5B \rightarrow 1s^2\ 2s^2\ 2p^1$

B^* $\quad$ [↑↓] [↑ | |] [↑] [↑ | ↑ |]

vacant p orbital

L.P = 3 − 3 = 0



50A°

Due to back-bonding, partial double bond character develops in B-F bond. So, the bond length becomes smaller.

$$\left[\begin{array}{c}\text{Bond length of}\\ \text{double bond}\end{array}\right] < \left[\begin{array}{c}\text{Bond length of}\\ \text{single bond}\end{array}\right]$$

eg.

Q) What is difference between Co-ordinate bond & Back Bond.

Co-ordinate Bond	Back Bond
1) This Bond is formed by overlapping of vacant orbital and lone pair orbital in different species.	1) This Bond is formed by overlapping of vacant orbital and lone pair orbital in same species.
2) This can be formed b/w orbitals of different energy level.	2) Back bond forms between orbitals of almost same energy level ↓ Due to this reason BF_3 exists but AlF_3 doesn't exist.

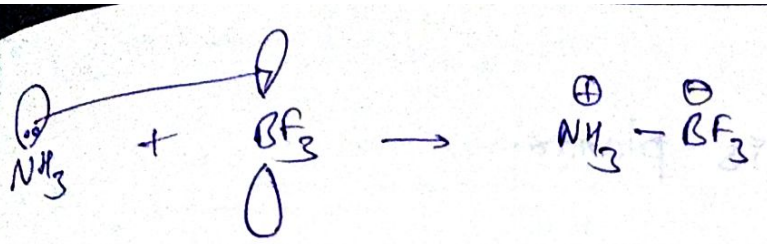
- If BF_3 has back bonding then why the below bond forms.

$$\ddot{N}H_3 + BF_3 \longrightarrow \overset{\oplus}{N}H_3 - \overset{\ominus}{B}F_3$$

Note :- If a compound has choice Back Bonding & Co-ordinate Bond. then,

Remember

$$\boxed{\text{Co-ordinate Bond} > \text{Back Bonding}}$$



Compare Lewis Acid Character:

Eg. BF_3 BCl_3 BBr_3 BI_3

→ BF_3 BCl_3 BBr_3 BI_3

↓ ↓ ↓ ↓

2p-2p 2p-3p 2p-4p 2p-5p

Back Bonding
↓ ke liy a
i) Vacant orbital
ii) Lone pair

overlapping of orbitals happens with almost same energy level → hence only BF_3 will show Back Bonding

$BF_3 > BCl_3 > BBr_3 > BI_3$ ... (Stability of Back Bond)

$BF_3 < BCl_3 < BBr_3 < BI_3$ ... (Lewis Acid)

IMPORTANT POINTS:

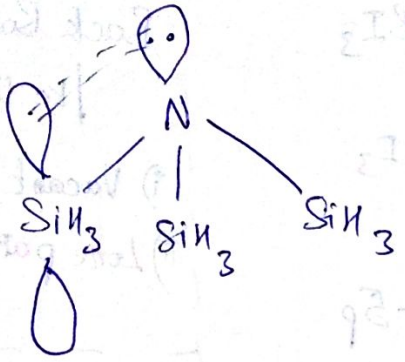
- Back Bond generally takes place b/w two bonded atoms. One bonded atom should have vacant orbital & other should have lone pair.

Back Bonding

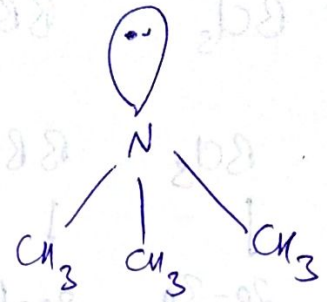
Eg. Tri-silyl Amine is planar

But

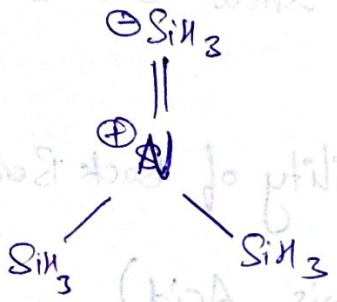
Trimethyl Amine is pyramidal



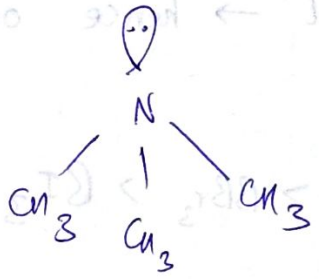
Tri Silyl amine



Tri-methyl Amine



Planar



Pyramidal

Note: Bond banane ke baad back Bond banthai. → To increase Stability.

→ In tri-silyl amine Si has vacant d-orbital
∴ Back Bonding occurs.

whereas, In trimethyl amine C does not have vacant orbital so Back Bonding doesn't occur.

Si → Belongs to 3rd period → ∴ It has vacant d-orbital

C → Belongs to 2nd period → ∴ It doesn't have any vacant orbital.

Banana Bonding
OR
Bonding in electron deficient Compound

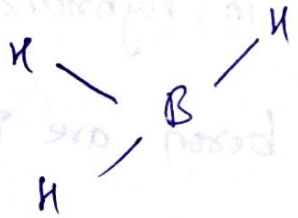
In electron deficient Compound we have :-

1) 3 centre 2 electron
2) 3 centre 4 electron

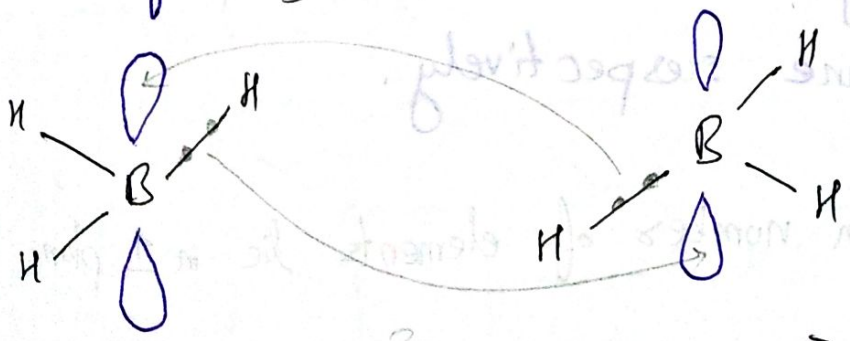
3 centre 2 electron

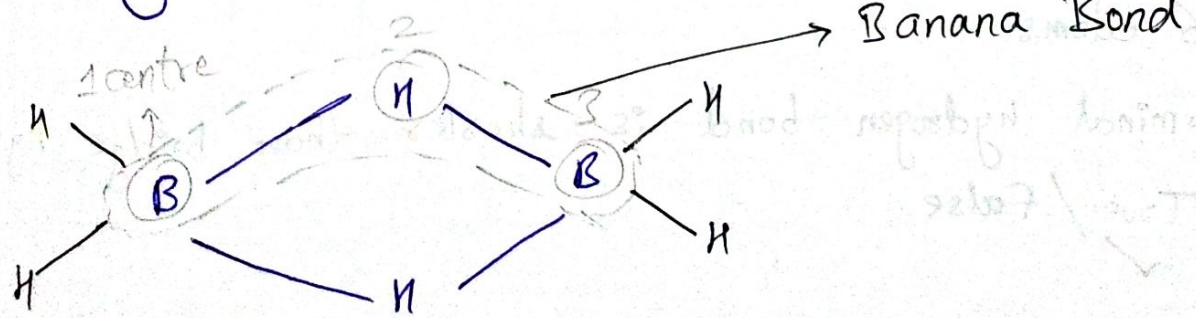
- BH_3 is a hypovalent molecule its octet is not complete.
- So, BH_3 tries to complete its octet by making dimers and polymers.

Draw Lewis Dot structure of BH_3

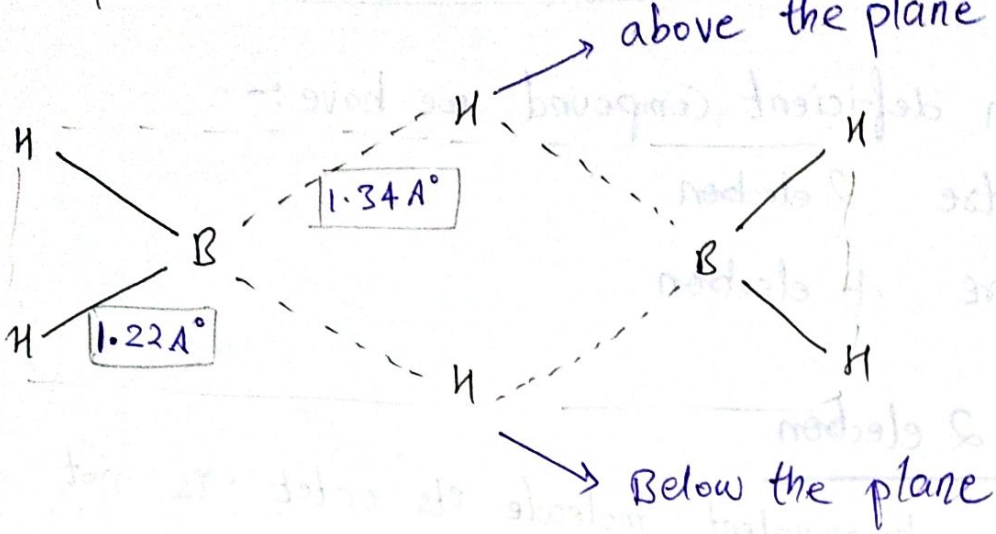


- Dimer of BH_3 (B_2H_6) [Di-Borane]





In Di-Boran (B_2H_6) two types of B-H Bonds are present.



VIP Points about B_2H_6

- Hybridization of B in BH_3 is sp^2 but in B_2H_6 it is sp^3
- Boron momentarily completes its octet.
- Vacant orbital of Boron takes part in Hybridization.
- 4 terminal hydrogen and two boron are in same plane.
- 2 Bridge hydrogen are above the plane & below the plane respectively.

Q) How many maximum number of elements lie in 1 plane in B_2H_6

→ 6 atoms

Q) Terminal hydrogen bond is shorter than bridge hydrogen

⇒ True / False

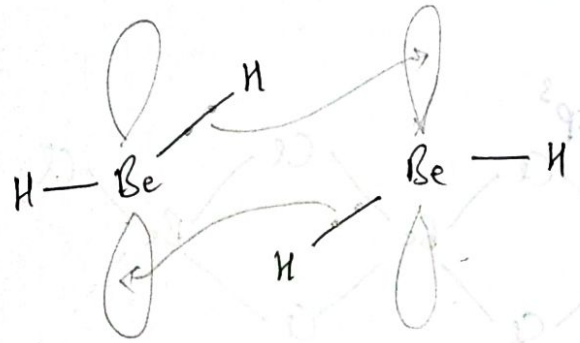
show the dimer & polymer of BeH_2

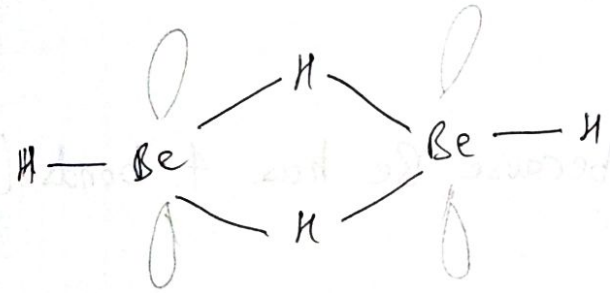
$^9Be \rightarrow 1s^2\ 2s^2$

↑↓			
↑	↑		

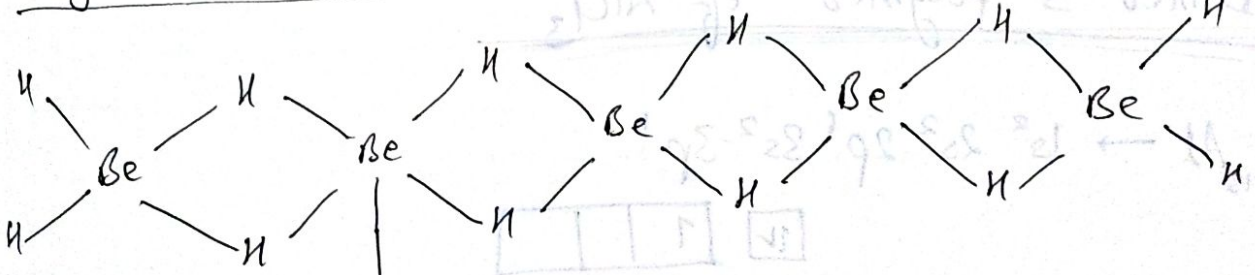
→ 2 vacant orbital ⇒ so forms dimer & polymer both

Dimer of BeH_2





Polymer of BeH_2



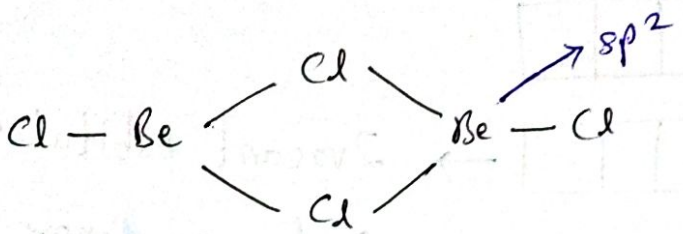
sp^3
(non - planar)
(3 centre $2e^-$)
(electron deficient)

$sp^2 \rightarrow$ planar

$sp^3 \rightarrow$
$sp^3d \rightarrow$ non - planar
$sp^3d^2 \rightarrow$ planar

3 centre 4 electron

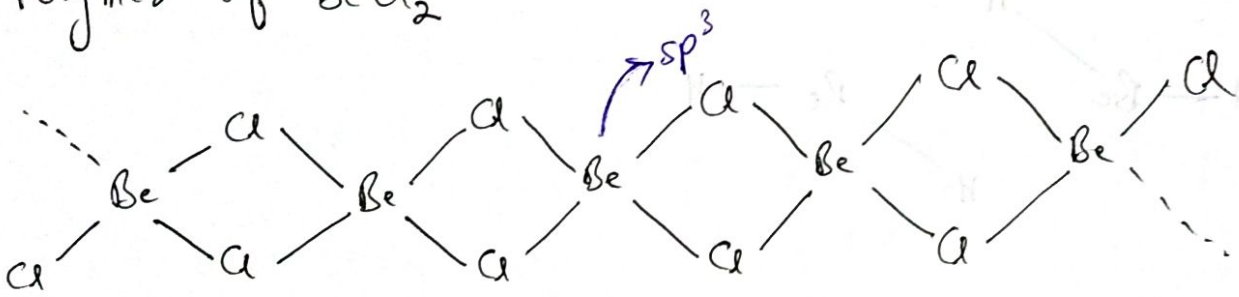
- Show dimer & polymer of $BeCl_2$



i) 3 centre - 4 electron Bond (Bridge atom has lone pair)

2) Electron deficient because Be has only 3 bonds ($6e^-$)

Polymer of $BeCl_2$



i) 3C - $4e^-$ bond

ii) Not electron deficient because Be has 4 bonds ($8e^-$)

iii) Non-planar

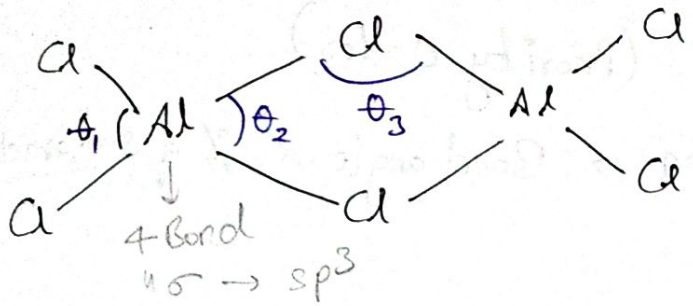
Dimer & polymer of $AlCl_3$

- $_{13}Al \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

	1↓	1		
Al^*	1	1	1	

one vacant orbital so only dimer is formed

Dimer of $AlCl_3$



Pair)

(6e⁻)

i) Non-planar (sp^3 is non-planar)

ii) Maximum no. of atoms in same plane = 6

iii) $\theta_1 > \theta_3 > \theta_2$

Comparison of Bond Angle

Follow only 4 steps: (Priority Order)

1. Calculate hybridization ⇒ Bond angle $\propto$ % of s-character

2. Bond angle $\propto \dfrac{1}{\text{Lone pair of central atom}}$

3. Bond angle $\propto$ Electronegativity of central atom

4. Bond angle $\propto \dfrac{1}{\text{Electronegativity of surrounding atom}}$

Percentage of s-character

$$\boxed{\% \text{ of s-character} = \frac{\text{no. of s-orbital}}{\text{Total no. of orbital}} \times 100}$$

1. Calculate hybridization.

Eg. Which of the following has largest Bond angle?

A) $BeCl_2$ B) BCl_3 C) CCl_4 d) PCl_5

→ Calculate hybridization

10

s-character

$BeCl_2$

v.e = 2 + 7×2
= 16

8) 16 (2
16
x

2 → sp

s% = $\frac{1}{2}$ × 100

s% = 50%

BCl_3

v.e = 3 + 3×7
= 24

8) 24 (3
24
(x)

3 → sp^2

s% = $\frac{1}{3}$ × 100

s% = 33%

CCl_4

v.e = 4 + 7×4
= 32

8) 32 (4
32
(x)

4 → sp^3

s% = $\frac{1}{4}$ × 100

s% = 25%

PCl_5

v.e = 5 + 8×7
= 40

8) 40 (5
40
(0)

5 → sp^3d

s% = $\frac{1}{5}$ × 100
= 20%

∴ Bond angle ∝ % of s-character

∴ $BeCl_2$ > BCl_3 > CCl_4 > PCl_5

Q which of the following has largest Bond angle?

→ NO_2^+

v.e = 5 + 6×2 − 1
= 16

8) 16 (2
16
(x)

2 → sp

% s = $\frac{1}{2}$ × 100 = 50%

NO_2^-

v.e = 5 + 6×2 + 1
= 18

8) 18 (2
16
2) 2 (1
2
(x)

3

3 → sp^2

% s = $\frac{1}{3}$ × 100 = 33%

∴ B.angle ∝ s% of character

∴ NO_2^+ > NO_2^-

?

Eg. which of the following has largest Bond angle?

NH_3	NH_4^+
$ve = 5 + 3\times1$	$v.e = 5 + 4 - 1$
$= 8$	$= 8$
$2\overline{)8}\,(4$ 8 $(\times)$	$2\overline{)8}\,(4$ 8 $(\times)$
$4 \rightarrow sp^3$	$4 \rightarrow sp^3$
$s\% = 25\%$	$s\% = 25\%$
$\therefore$ L.P on C.A $= 1$	L.P on C.A $= 4-4 = 0$
$\because$ Bond angle $\propto \dfrac{1}{\text{L.P on C.A}}$	$\because$ Bond angle $\propto \dfrac{1}{\text{L.P on C.A}}$

$\therefore$ $\boxed{NH_4^+ > NH_3}$ → Ans

Eg. which of the following has largest Bond angle?

SO_3

$v.e = 6 + 3 \times 6$

$= 24$

$8 \overline{)24} (3$

$\underline{24}$

$(\times)$

$3 \rightarrow sp^2$

$L.Pe = 3 - 3 = 0$

SO_2

$v.e = 6 + 6 \times 2 = 18$

$8 \overline{)18} (2$

$\underline{16}$

$2 \overline{)2} (1$

$\underline{2}$

$(\times)$

> 3

$3 \Rightarrow sp^2$

$L.Pe = 3 - 2 = 1$

$\therefore$ Bond angle $\propto \dfrac{1}{L.P \text{ on } C.A}$

$\therefore SO_3 > SO_2$

Eg. which of the foll has largest Bond angle?

CH_4	NH_3	H_2O
$\rightarrow sp^3$	sp^3	sp^3
$L.P = 4 - 4 = 0$	$L.P = 4 - 3 = 1$	$L.P = 4 - 2 = 2$

$\because$ Bond angle $\propto \dfrac{1}{L.P \text{ on } C.A}$

$\therefore CH_4 > NH_3 > H_2O$

Eg. which of the foll has largest Bond angle?

PI_3	AsI_3	SbI_3
→ sp^3	sp^3	sp^3
$L.P = 4-3 = \underline{1}$	$L.P = 4-3 = 1$	$L.P = 4-3 = \underline{1}$

Step 3: Bond angle $\propto$ Electronegativity of central atom.

$\therefore$ E.N order = P > As > Sb

$\therefore$ Bond angle = $PI_3 > AsI_3 > SbI_3$

Eg. which of the following has largest Bond angle?

NH_3 PH_3 AsH_3 SbH_3

→ • Hybridization is same
• Lone pair same
• EN of central atom ✓

$$\boxed{NH_3 > PH_3 > AsH_3 > SbH_3}$$

Eg. which of the following has largest Bond angle?

H_2O H_2S H_2Se H_2Te

→ • Hybridization same

• Lone pair same

• EN of central atom ✓

$$\boxed{\text{Bond angle} \propto \text{E.N of C.A}}$$

$$\boxed{\therefore H_2O > H_2S > H_2Se > H_2Te}$$

Eg. which of the following has largest Bond angle?

SF_2 SCl_2 SBr_2 SI_2

→ • Hybridization is same

• L.P. same

• E.N of central atom same

then, $$\boxed{\text{Bond angle} \propto \frac{1}{\text{E.N of surrounding atom}}}$$

$\therefore$ Bond angle :- $SF_2 < SCl_2 < SBr_2 < SI_2$

Q) CoF_2 $CoCl_2$ $CoBr_2$ CoI_2

→ • Hybridization same
• Lone pair same
• EN of central atom same
• EN of surrounding atom ✓

E.N order : $F > Cl > Br > I_2$

∴ Bond angle ⇒ $CoF_2 < CoCl_2 < CoBr_2 < CoI_2$

DRAGO'S RULE :

Drago's Rule is applied when :

- Central atom should be from G15, G16, G17
- C.A should be 3rd period or higher.
- C.A should have lone pair
- Electronegativity of surrounding atom should be less than equal to 2.5

Drago's Rule :- If the central atom of a molecule belongs to 15th or 16th group & is in 3rd period or below, and is bonded to a substituent having E.N ≤ 2.5, then L.P present in s-orbital doesn't participate in hybridization and bonds are formed by pure p-orbitals. As pure p-orbitals are at 90° to each other, angle b/w bond pairs are found to be close to 90°. L.P of s-orbital lies in stereo-chemically inactive s-orbital

2p

NH_3 107°	>	$\ddot{P}H_3$ 92°	>	$\ddot{As}H_3$ 91°	>	$\ddot{Sb}H_3$ 90°
H_2O 104°	>	$H_2\ddot{S}$ 92°	>	$H_2\ddot{Se}$ 91°	>	$H_2\ddot{Te}$ 90°

$H \rightarrow E.N \rightarrow 2.1 \leq 2.5$

Example of Drago's rule

Eg. $\ddot{P}H_3$ > $\ddot{As}H_3$ > $\ddot{Sb}H_3$

92° 91° 90°

Eg. $\ddot{S}H_2$ > $\ddot{Se}H_2$ > $\ddot{Te}H_2$

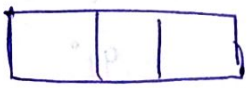
92° 91° 90°

Eg. $\ddot{Cl}-H$ > $\ddot{Br}-H$ > $\ddot{I}-H$

92° 91° 90°

Why Bond angle is near to 90°

- Pure orbitals will participate in Bonding
- means bond angle will be always near 90°

Pure orbitals → 

P_x P_y P_z

sp^3, sp^2

\# Eg. Compare Bond Angle. (NEET/AIIMS)

	NH_3	NF_3	PF_3	PH_3
→	sp^3	sp^3	sp^3	(Drago's Rule) ⇓
L.Pe :	1	1	1	(So angle 90°)

Now, compare NH_3 NF_3 PF_3

E.N of N > P

∴ Bond angle ∝ E.N of central atom

E.N of F > H

∴ Bond angle $\propto \frac{1}{\text{E.N of surrounding atom}}$

$$\boxed{\therefore NH_3 > NF_3 > PF_3 > PH_3}$$

Bond length

Comparison of Bond length :

1) Bond length $\propto \frac{1}{\Delta \text{ Electro negativity}}$

2) Bond length $\propto \frac{1}{\text{Bond order}}$

3) Resonance → Due to resonance bond length decreases.

4) Due to Back-Bonding bond length decreases.

Eg. Compare Bond length

i) $C-N$, $C=N$, $\overset{\ominus}{C}\equiv N^{\oplus}$

→ Bond length $\propto \dfrac{1}{\text{Bond order}}$

$\therefore$ B.L ⇒ $\boxed{C-N > C=N > \overset{\ominus}{C}\equiv N^{\oplus}}$

ii) $C-C$, $C-N$, $C-O$

→ Bond length $\propto \dfrac{1}{\Delta E.N}$

E.N. : $C-C \Rightarrow 2.5-2.5 = 0$

$C-N \Rightarrow 2.5-3 = 0.5$

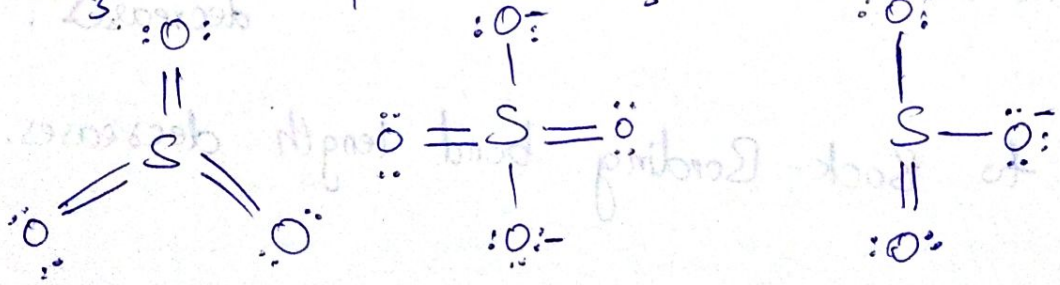
$C-O \Rightarrow 2.5-3.5 \Rightarrow 1$

$\therefore$ B.L ⇒ $\boxed{C-C > C-N > C-O}$

Eg. Compare Bond length :

SO_3 $\quad SO_4^{-2}$ $\quad SO_3^{-2}$

⇒



$B.O = 2$

$B.O = 2 + \frac{(-2)}{4}$

$B.O = 1.5$

$B.O = 2 + \frac{(-2)}{3}$

$B.O = 1.4$

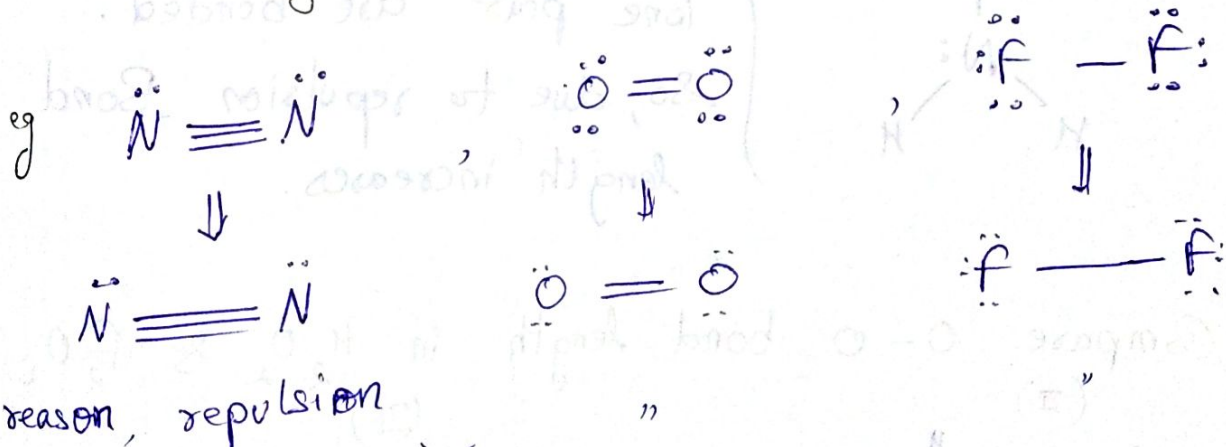
$B.O = 2$

$\therefore$ Bond length $\propto \frac{1}{\text{Bond order}}$

$\therefore B.L \Rightarrow SO_3^{-2} > SO_4^{-2} > SO_3$

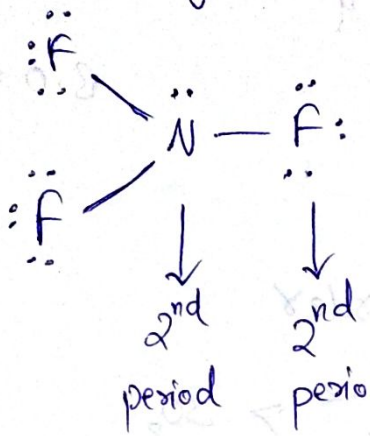
VIP Points

- If two 2nd period elements having lone pair are bonded then due to repulsion, Bond length increases.

eg 

reason, repulsion between lone pair

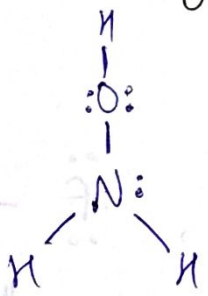
☆ Eg N-F bond length in NF_3 is higher than expected, why?

→ 

$Q = 5 + 7 \times 3$
$= 26e^-$
$B.Pe^- = 6e^-$
$L.Pe^- = 26 - 6 = 20e^-$

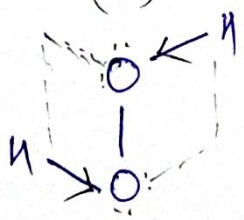
So, both are 2nd period & have lone pair thus, due to repulsion bond length increases

Eg. N-O Bond length in NH_2-OH is higher than expected why?

⇒ 

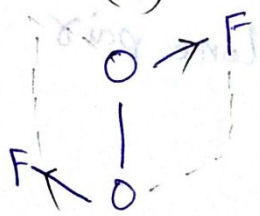
2nd period elements with lone pair are bonded. So, due to repulsion Bond length increases.

Eg. Compare O-O bond length in H_2O_2 & F_2O_2

(I) 

Open Book structure

E.N : O > H

(II) 

Open Book structure

E.N :- F > O

xpected? why?

electron density will increase on oxygen in first st (I) due to E.N: O > H & in second structure it will not. because E.N of F > O

∴ Bond length : $H_2O_2 > F_2O_2$

oe–

pair

ases

BOND ENERGY

i) B.E ∝ Δ Electronegativity

ii) B.E ∝ B.O

3) $B.E \propto \frac{1}{\text{Size of atom}}$

Compare Bond Energy

H–F H–Cl H–Br H–I

→ H–F, H–Cl, H–Br, H–I

2.1 – 4 = 1.9

2.1 – 3 = 0.9

2.1 – 2.8 = 0.7

2.1 – 2.5 = 0.4

∴ Bond energy ∝ Difference in E.N

$\boxed{\therefore H-F > H-Cl > H-Br > H-I}$

eg. Compare bond energy

~~Si-Si~~ Si-Si P-P S-S Cl-Cl

→ Size : Si > P > S > Cl

Bond energy $\propto \frac{1}{\text{Size of atom}}$

$\boxed{\therefore \text{Si-Si} < \text{P-P} < \text{S-S} < \text{Cl-Cl}}$

eg. Compare Bond energy

C-C $\ddot{N}-\ddot{N}$ $\ddot{O}-\ddot{O}$ $:\ddot{F}-\ddot{F}:$

→ No lone pair repulsion in carbon ← C-C > Si-Si

(Size↑ ⇒ B.E↑)

2nd Period element having lone pair repels each other & so, bond length increases ← N-N < P-P

O-O < S-S

F-F < Cl-Cl

(Size↑ ⇒ B.E↓)

Exceptio

Exception

F_2

→ Excep

Bo

Experim

Dip

• Dem

Prop

• Ve

eg.

Exception

Exception of B.E in Halogen group:

F_2 Cl_2 Br_2 I_2

→ Exceptation: $F_2 > Cl_2 > Br_2 > I_2$

But

Experimentally: $\boxed{Cl_2 > Br_2 > F_2 > I_2}$ (Bond energy)

Dipole Moment

Di – pole
↓ ↓
Two ⊕ & ⊖

- Denoted by $\vec{\mu}$

Properties of Dipole Moment

- Vector quantity

$$\boxed{\vec{\mu} = |q| \times d}$$

$|q|$ → magnitude of charge
d → distance Btw two atoms / Bond length

eg. $\overset{\delta+}{H}$ —— $\overset{\delta-}{Cl}$
2.1 3

Direction

i) Less EN to More E.N.

ii) In case of lone pair → Atom to lone pair

eg. $\overset{\delta+}{H}$ — $\overset{\delta-}{Cl}$

2.1 3

→

Dipole moment in Molecules :

(1) AB_2 (2) AB_3 (3) AB_4 (4) AB_5 (5) AB_6

(6) AB_7

I) AB_2 type Molecules

eg. $BeCl_2$

Cl — Be — Cl

μ μ

opposite vectors of same magnitude cancel each other

$\boxed{\mu_{net} = 0}$

∴ Non-polar

eg. CO_2

O = C = O

μ μ

opposite vectors of same magnitude cancel each other

$\boxed{\mu_{net} = 0}$

∴ Non-polar

eg C_2H_2

H — C ≡ C — H

opposite vectors of same magnitude cancel each other

$\boxed{\mu_{net} = 0}$

∴ Non-polar

Note :

E.N:

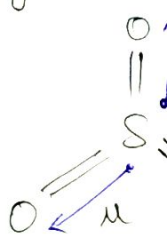
$\boxed{E.N \propto}$

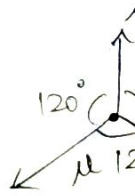
$\mu_{net} =$

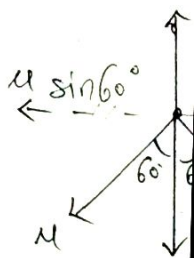
$\mu_{net} \neq$

II) AB_3

eg. SO







Note : $sp > sp^2 > sp^3$

E.N: 2.9 2.7 2.5

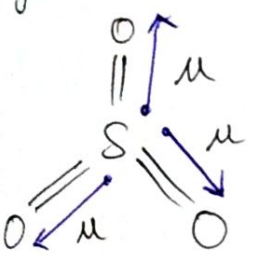
$\boxed{E.N \propto \%\text{ of } s\text{- character}}$

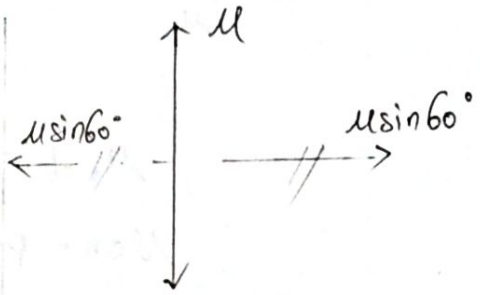
$\mu_{net} = 0$ (Non- polar)

$\mu_{net} \neq 0$ (Polar)

II) AB_3 type molecules

eg. SO_3



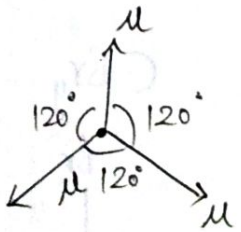


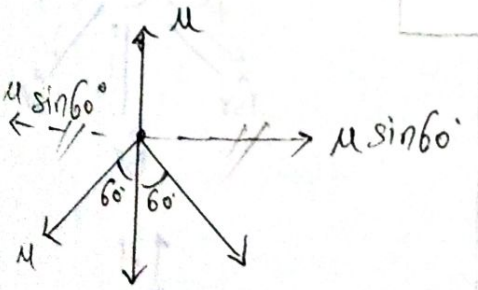
$\mu\cos60° + \mu\cos60°$

$\mu\left(\frac{1}{2}\right) + \mu\left(\frac{1}{2}\right)$

$\frac{\mu}{2} + \frac{\mu}{2}$

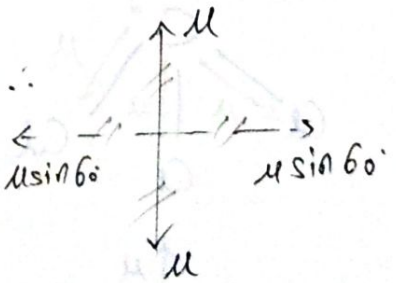
μ





$\mu\cos60°$
$+$
$\mu\cos60°$

$\therefore$



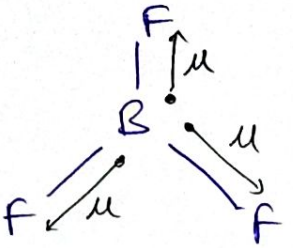
... opposite vectors of same magnitude cancel each other

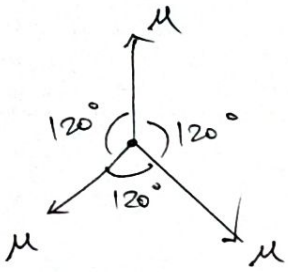
$\boxed{\mu_{net} = 0}$

Non-polar

Similiar eg BF_3 , BCl_3 as SO_3

BF_3

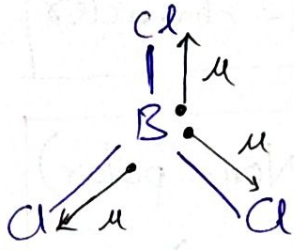


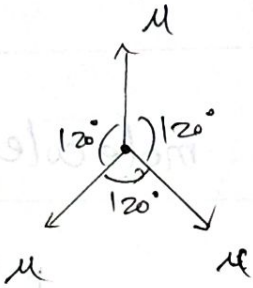


$\boxed{\mu_{net} = 0}$

Non- polar

BCl_3

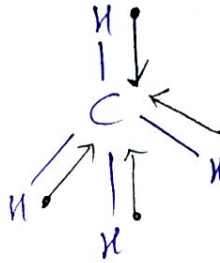




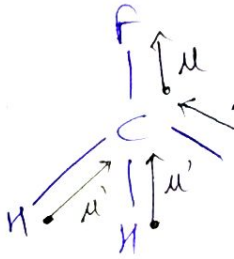
$\boxed{\mu_{net} = 0}$

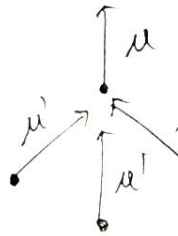
Non- polar

eg. CH_4



eg. CH_3F



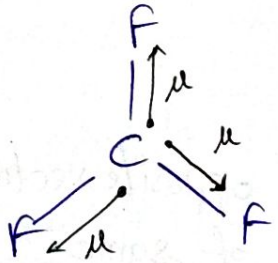


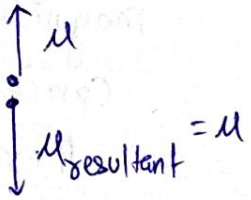
$\mu_{net} \neq$

∴ Polar

III) AB_4 type molecules

eg. CF_4

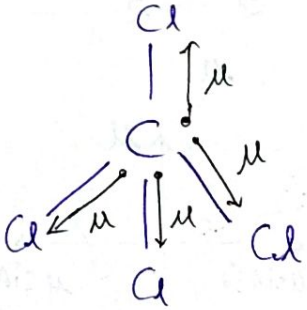


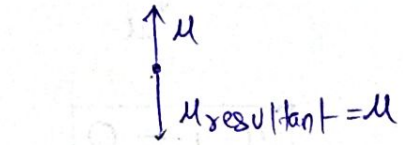


$\boxed{\mu_{net} = 0}$

(non- polar)

CCl_4

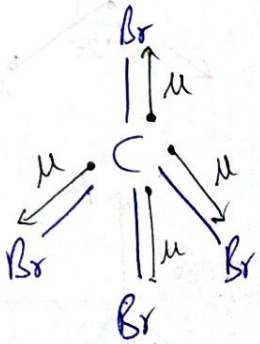


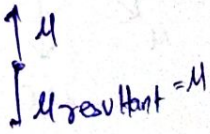


$\boxed{\mu_{net} = 0}$

(Non- polar)

CBr_4

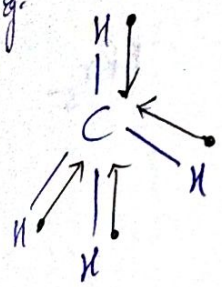


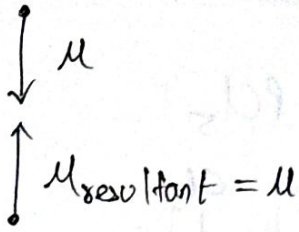


$\boxed{\mu_{net} = 0}$

(Non- polar)

eg. CH_4



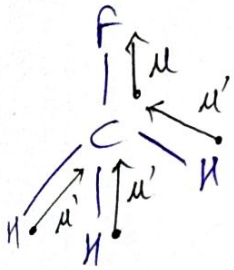


Equal & opposite vectors cancel each other

$\boxed{\therefore \mu_{net} = 0}$

(Non-polar)

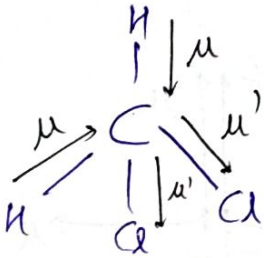
eg. CH_3F

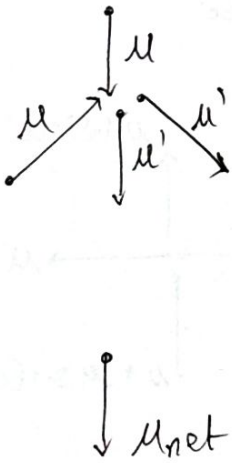


$\mu_{net} \neq 0$

∴ polar molecule

CH_2Cl_2

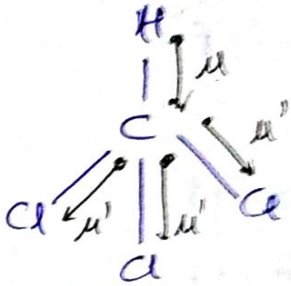


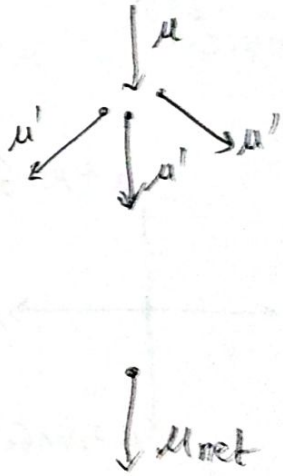


$\mu_{net} \neq 0$

∴ polar molecule

$CHCl_3$



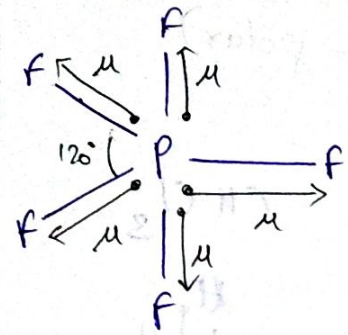


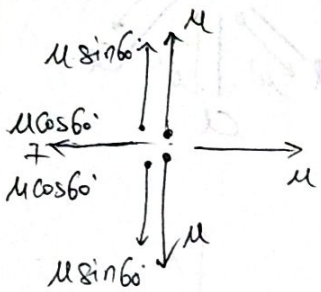
$\mu_{net} \neq 0$

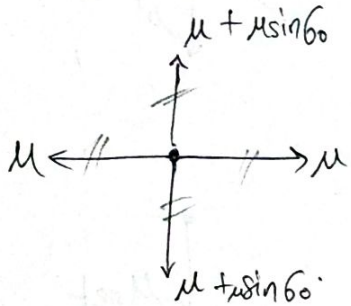
∴ Polar mol.

IV) AB_5 type molecules

eg. PF_5



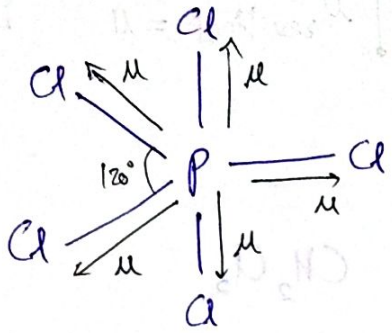


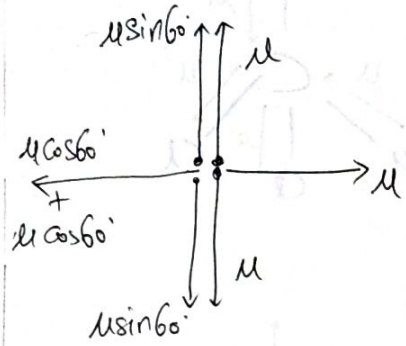


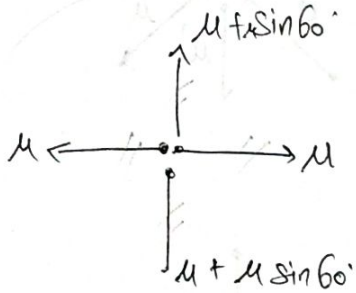
$\boxed{\mu_{net} = 0}$

∴ Non-polar

PCl_5







$\boxed{\mu_{net} = 0}$

∴ Non-polar

PBr_5

$\boxed{\mu_{net} = 0}$

∴ Non-polar

V) AB_6

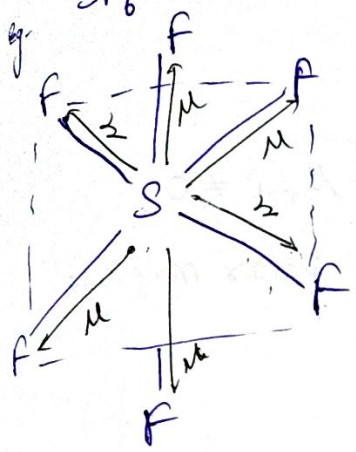
eg. SF

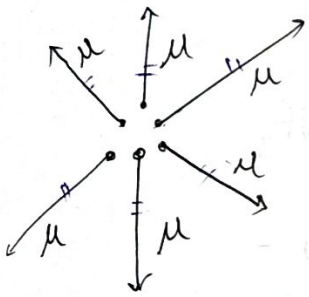
$\boxed{\mu}$

∴ No

I) AB_6 type molecules

eg. SF_6



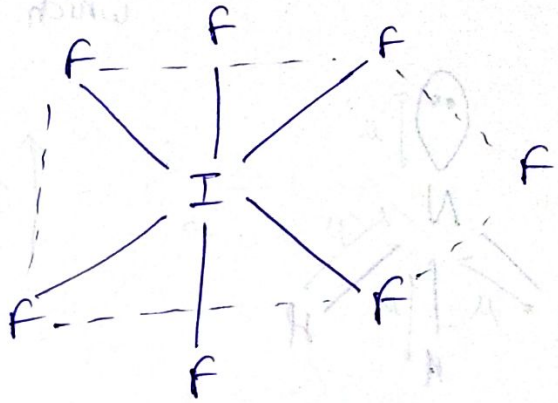


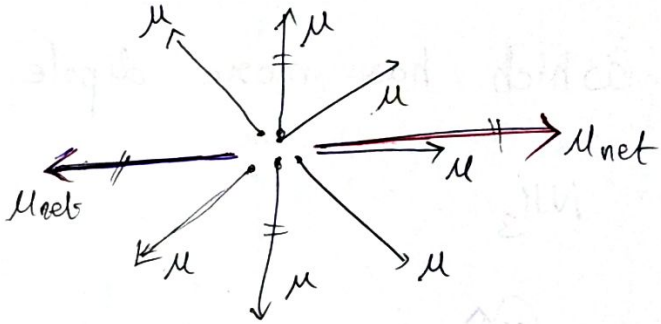
$\boxed{\mu_{net} = 0}$

∴ Non-polar

AB_7 type molecules

eg IF_7



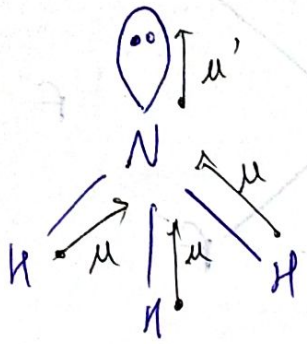


$\boxed{\mu_{net} = 0}$

∴ Non-polar

Example of molecule having lone pair

NH_3 find: Net dipole moment will point in which direction

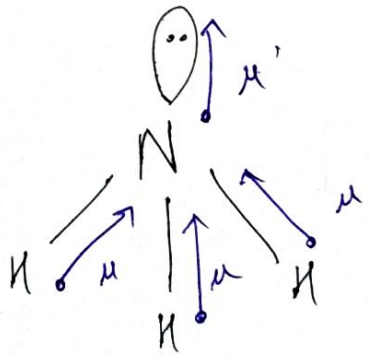


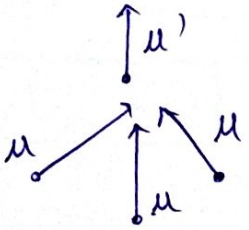
$\Rightarrow$ $3\mu + \mu'$ $\Rightarrow$ $\boxed{\mu_{net} \neq 0}$

Polar molecule

Eg which has more dipole moment?

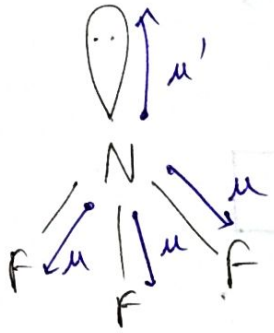
NH_3

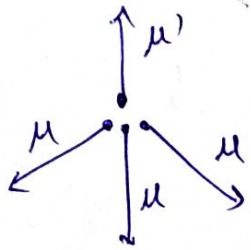




All vectors will add & μ_{net} will be in one direction. (Result will be addition of four vectors)

NF_3



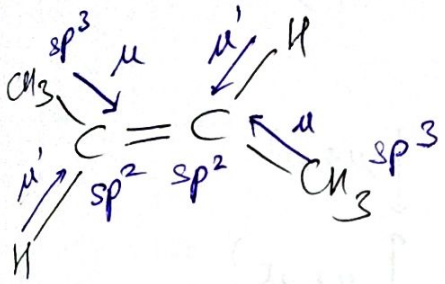


Resultant vector will be (addition of 3 vectors − 1 vector)

$\therefore NH_3 > NF_3$ (dipole moment)

which has more dipole moment.

eg.



(trans)

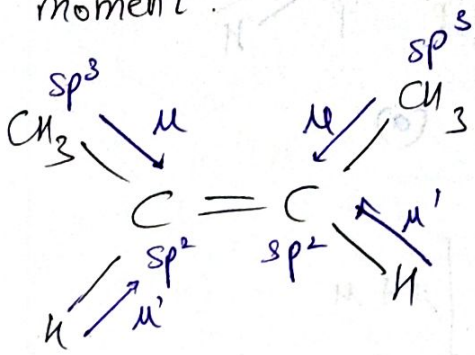
$\downarrow \mu + \mu'$

$\uparrow \mu' + \mu$

equal & opposite vectors cancel each other.

$\boxed{\mu_{net} = 0}$

$\therefore$ Non-polar



(cis)

$\downarrow \mu + \mu$

$\uparrow \mu' + \mu'$

opposite ✓

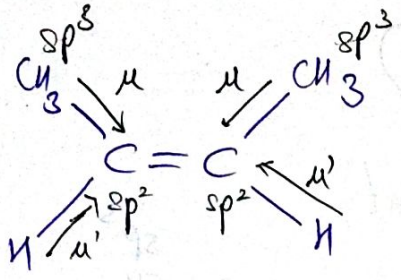
equal ✗

$\therefore$ Not cancel each other

$\boxed{\therefore \mu_{net} \neq 0}$

polar

Eg.



(a)

$\mu + \mu$

$\mu' + \mu'$

opposite ✓

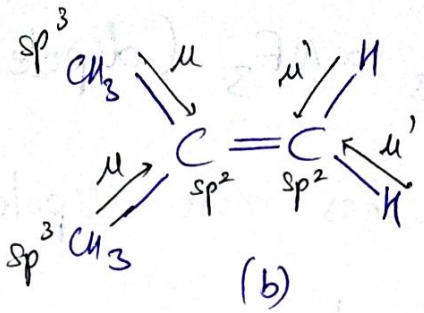
equal ✗

∴ Not cancel each other

$\boxed{\mu_{net} \neq 0}$

polar

$(a > b) \rightarrow$ Dipole



(b)

$\mu + \mu'$

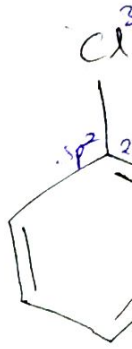
$\mu + \mu'$

opposite & equal vectors cancel out each other.

$\mu_{net} = 0$

Non-polar

Eg



(a

~~Equal~~

$\boxed{\mu_{net}}$

(polar

(a

Eg

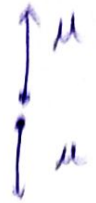
(a)

(b)

~~Equal~~

$\boxed{\mu_{net} \neq 0}$

(polar)

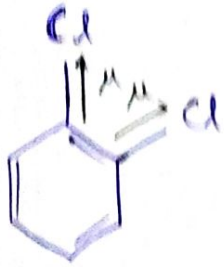


Equal & opposite cancel each other

$\boxed{\mu_{net} = 0}$

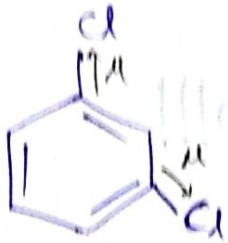
(Non-polar)

(a > b) → Dipole moment



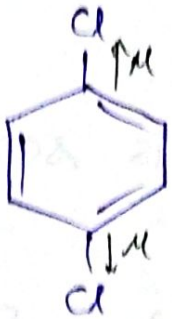
(a)

$\theta = 60°$
(ortho)



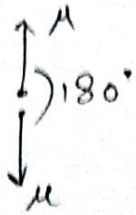
(b)

$\theta = 120°$
(meta)



(c)

$\theta = 180°$
(para)



$$\mu_{net} = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2\cos\theta}$$

$$= \sqrt{\mu^2 + \mu^2 + 2\mu^2\cos 60^\circ}$$

$$= \sqrt{2\mu^2 + \cancel{2}\mu^2 \tfrac{1}{\cancel{2}}}$$

$$= \sqrt{2\mu^2 + \mu^2}$$

$$= \sqrt{3\mu^2}$$

$$\boxed{\mu_{net} = \sqrt{3}\,\mu}$$

$$\mu_{net} = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2\cos 120^\circ}$$

$$= \sqrt{\mu^2 + \mu^2 + \cancel{2}\mu^2\left(-\tfrac{1}{\cancel{2}}\right)}$$

$$= \sqrt{2\mu^2 - \mu^2}$$

$$= \sqrt{\mu^2}$$

$$\boxed{\mu_{net} = \mu}$$

$$\mu_{net} = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2\cos 180^\circ}$$

$$= \sqrt{2\mu^2 + 2\mu^2(-1)}$$

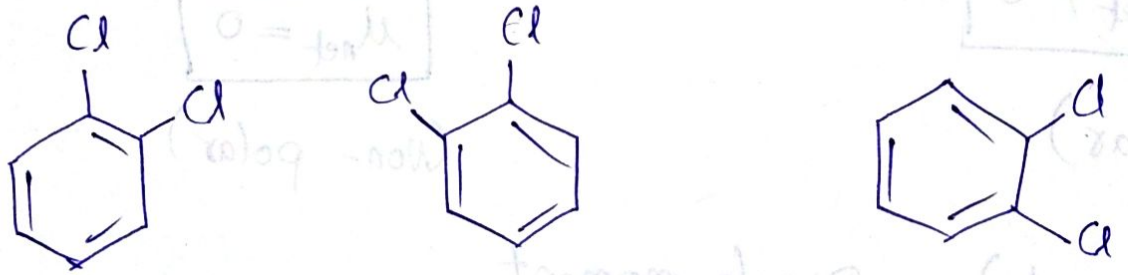
$$= \sqrt{2\mu^2 - 2\mu^2}$$

$$= 0$$

$$\boxed{\mu_{net} = 0}$$

a > b > c → Dipole moment

Eg.



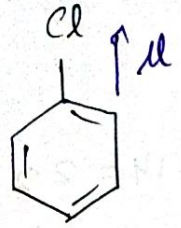
$\theta = 60^\circ$ $\theta = 60^\circ$ $\theta = 60^\circ$

μ → is same mitron!!!

Note :- sp^3 → Non-planar

sp^2 → planar → Ek plane mai hota hai.

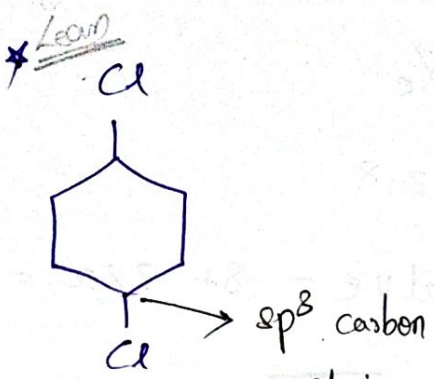
Eg which compound is non-polar?



(a)

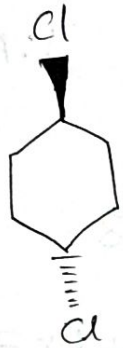
$\boxed{\mu_{net} \neq 0}$

(polar)



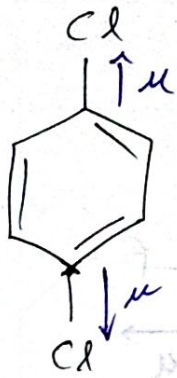
so, it is not planar

⇓



$\boxed{\mu \neq 0}$

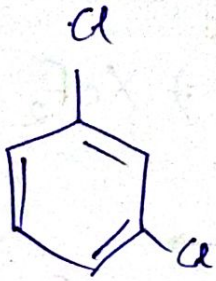
(polar)



Equal & opp: vectors cancel each other.

$\boxed{\mu = 0}$

Non-polar

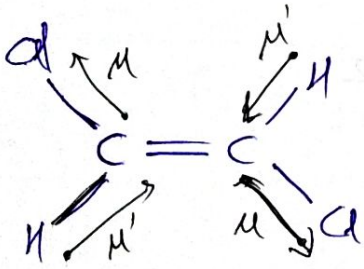


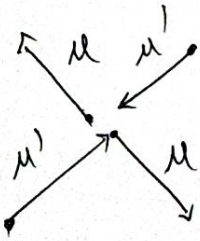
$\mu_{net} = \mu$

$\boxed{\mu \neq 0}$

(polar)

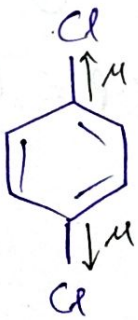
Q) which compound is non-polar?

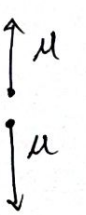




$\mu = 0$

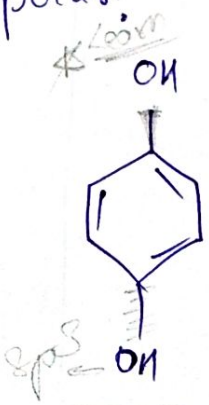
(NP)





$\boxed{\mu = 0}$

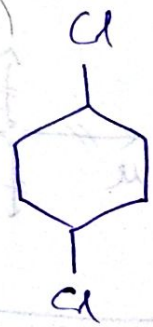
(Non-polar)



$\boxed{\text{Polar}}$

$\mu_{net} \neq 0$

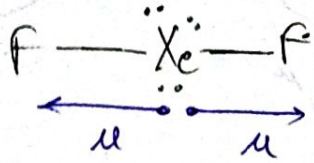
$\boxed{\text{Non-planar}}$



Eg. Calculate Dipole moment & check polarity:

i) XeF_2

→



$\boxed{\mu_{net} = 0}$

∴ Non-polar

$_{58}Xe$

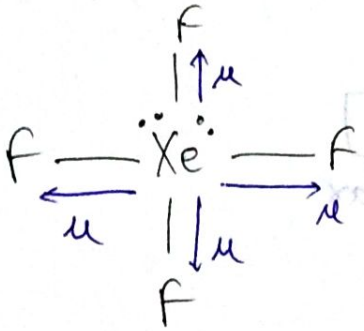
v.e = 8

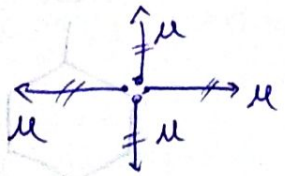
Total v.e = 8 + 7×2 = 8 + 14 = 22

$$\begin{array}{r} 8\overline{)22}(2 \\ 16 \\ \hline 2)\ 6\ (3 \\ 6 \\ \hline (\times) \end{array} \Big\rangle\ 5 \rightarrow sp^3d$$

$5 \rightarrow sp^3d$ L.Pe = 5 − 2 = $\boxed{3}$

ii) XeF_4





$\boxed{\mu_{net} = 0}$

∴ Non-polar

$_{58}Xe$ v.e = 8

Total v.e = 8 + 4×7 = 8 + 28 = 36

$$\begin{array}{r} 8\overline{)36}(4 \\ 32 \\ \hline 2)\ 4\ (2 \\ 4 \\ \hline (\times) \end{array} \Big\rangle\ 6 \rightarrow sp^3d^2$$

L.P = 6 − 4 = 2

12

8+14 = 22

d

-2 = $\boxed{3}$

28 = 36

2

Ex) $XeF_5^{\ominus}$

Xe

V.e = 8 + 5×7 + 1
= 8 + 35 + 1 = 44

8) 44 (5
40
2) 4 (2
4
(×)

7 → sp^3d^3

L.Pe = 7 − 5 = 2

Eg. which has more dipole moment?

1) NH_3 , H_2O , HF

→ $\mu = |q| \times d$

$q\uparrow \Rightarrow \mu\uparrow$

Jiska E.N ↑ ⇒ uska q↑

∴ F > O > N

$\boxed{\therefore HF > H_2O > NH_3}$

2) H_2O , H_2S , H_2Se , H_2Te

→ E.N → O > S > Se > Te

$\boxed{\therefore H_2O > H_2S > H_2Se > H_2Te}$ → Dipole moment

Exception In Dipole Moment

Q. Which has more dipole moment ?

CH_3F CH_3Cl CH_3Br CH_3I

$$\boxed{CH_3Cl > CH_3F > CH_3Br > CH_3I}$$

↓

net product of "q" & "d" is more for CH_3Cl than CH_3F (experimental fact)

Vander - wall force of attraction

Bond

- Strong Bond (200 - 400 kJ/mol)
  - ↓
  - Ionic + Covalent + Co-ordinate
- Weak Bond (2 - 4 kJ/mol)
  - Vander Wall forces (2-8 kJ/mol)
  - Hydrogen Bond (8 - 40 kJ/mol)

It is a very weak bond.

- This force of attraction acts between $\boxed{\text{molecules.}}$

$$\boxed{\text{Vander-wall force} \propto \text{Molecular weight}}$$

$$\boxed{\text{V.F} \propto \text{Molecular weight} \propto \text{Boiling point}}$$

Eg. F_2, Cl_2 are gases, whereas Br_2 is liquid & I_2 is solid. Explain.

→ F_2, Cl_2, Br_2, I_2 ↓ Mol. wt ↑ ⇒ Vander. W.F ↑

⇒ F_2, Cl_2 → gases
Br_2 → liquid
I_2 → ~~gase~~ solid

Secondary Bonds

Vander-wall force of attraction

Hydrogen-Bond

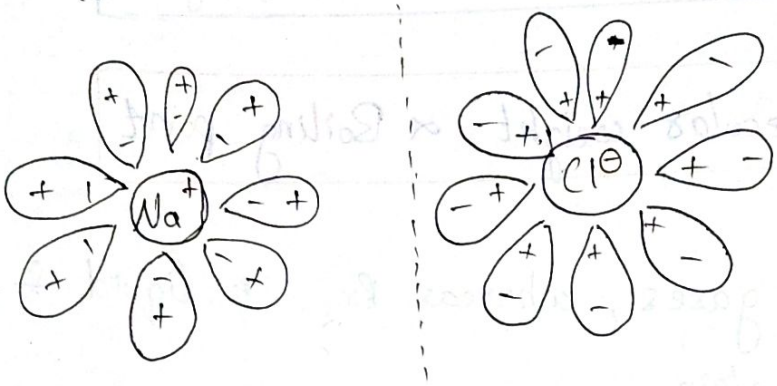
5 types of Vander-wall force

1. Ion - Dipole
2. Dipole - Dipole
3. Ion induced dipole
4. Dipole induced dipole
5. Instantaneous dipole induced dipole.

(I) Ion - Dipole

- This occurs between an ion and a polar molecule.

$$NaCl_{(s)} + H_2O \longrightarrow Na^+_{(aq)} + Cl^{\ominus}_{(aq)}$$



(II) Dipole - Dipole

- Occurs between two polar molecule.

HCl

HBr

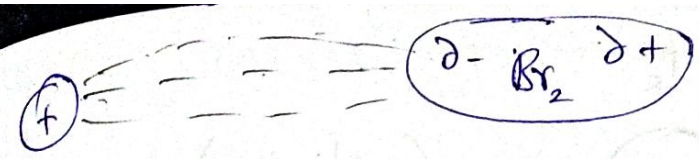
(III) Ion - Induced dipole

- Occurs between ion & non polar molecule
- Due to ion, polarity generates in non-polar molecule & then attraction occurs.

Na^+ & Br_2

(initially)

olecule.



Na

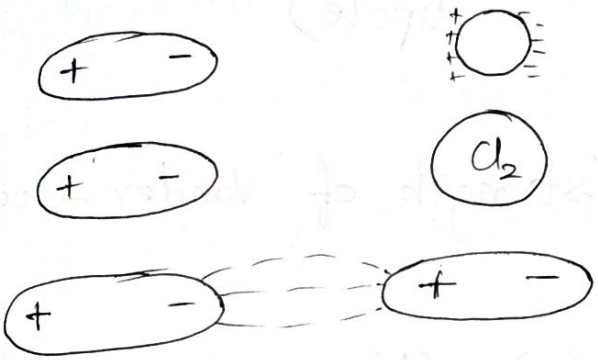
ion

(induced dipole by ion)

II) Dipole - Induced Dipole

- Occurs b/w dipole & non-polar molecule.

HCl & Cl_2



initially

(induced dipole by another dipole)

I) Instantaneous dipole Induced Dipole

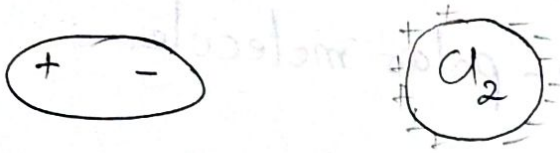
- Occurs between two non-polar molecules.
- One non-polar molecule dipole is instaneously created which induces dipole in other non-polar molecule.

Also known as London force or Dispersion force

eg. Cl_2 Cl_2 (initially)

↓

instantaneously dipole is created



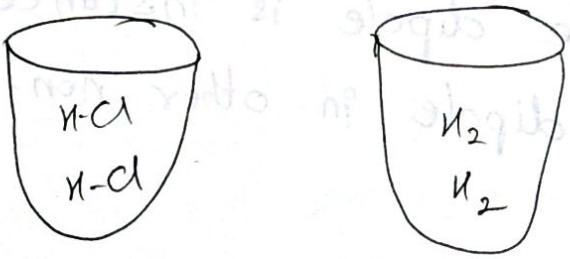
(induced dipole by another dipole)

- Increasing order of strength of Vander-wall attraction

(5) < (4) < (3) < (2) < (1)

Eg. Boiling point of H-Cl is more than H_2. Explain.

→



H-Cl	H_2
↓	↓
polar	Non-polar
↓	↓
Dipole	Instantaneous dipole induced dipole
Dipole attraction	

Eg. which has max. Boiling point?

A) He B) Ar C) Kr D) Xe

→ molecular wt. g ∝ Vanderwall force ∝ Boiling point

Eg. Compare B.P

H_2 Cl_2 N_2 Br_2 C_2 P_2

→ Instantaneous dipole induced dipole

Molecular wt ∝ Boiling point

$H_2 < C_2 < N_2 < P_2 < Cl_2 < Br_2$

Eg. Compare V.W.F

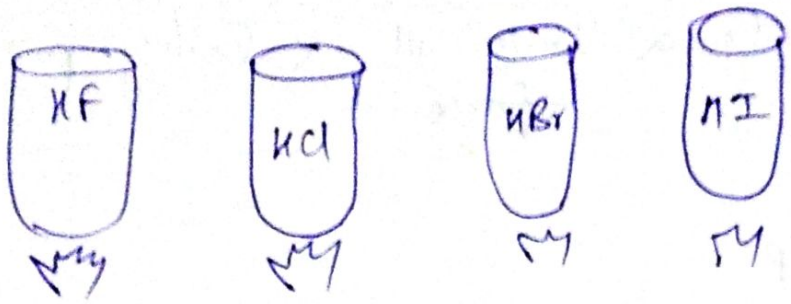
H_2 & H_2	Na^+ & NaCl	HCl & N_2
(a)	(b)	(c)
→ Instantaneous dipole induced dipole	Ion - dipole	Dipole induced Dipole

$\boxed{b > c > a}$

Q) Compare Boiling point

HF HCl HBr HI

→



Dipole-Dipole " " "

⇓

{H-Bonding}

Boiling . P ∝ Vander. W. F ∝ Molecular wt.

$$\boxed{HCl < HBr < HI < HF}$$

Eg. Compare Boiling point:

H_2O H_2S H_2Se H_2Te

H_2O ⇓ Hydrogen Bonding

H_2S, H_2Se, H_2Te: dipole - dipole

$$\boxed{\text{Boiling . P} \propto \text{Vander W. F} \propto \text{Molecular weight}}$$

$\therefore$ $H_2O > H_2Te > H_2Se > H_2S$

Hydrogen Bond → (only with 3 atoms)

A —— H = = = = = :B
(F/O/N) (F/O/N)

Hydrogen Bond wale FON karte hain.

- A should be F/O/N
- Hydrogen Bonding is electrostatic attraction.
- A should be attached to hydrogen.
- B should have Lone pair & B should also be F/O/N.

A —— H = = = = :B

↓ Covalent Bond (200 - 400 kJ/mol)

↓ Hydrogen Bond (8 - 40 kJ/mol)

Hydrogen Bond ki Mechanism

δ- δ+
A —— H :B

E.N atom

- As we know "A" is highly electronegative atom so, it will attract electrons towards itself.

$$\overset{\delta-}{A} \text{———} \overset{\delta+}{H} \text{- - - -} \ddot{B}$$

H becomes electron deficient. So, it forms bond with electronegative atom which has Lone pair.

$$\overset{\delta-}{A} \text{———} \overset{\delta+}{H} \text{- - - - -} \ddot{B}$$

This electrostatic attraction b/w hydrogen & B (with lone pair) is called hydrogen bond.

STRENGTH OF HYDROGEN BOND

$$A \text{———} H \text{- - - - - -} \ddot{B}$$

i) Strength of H-Bond $\propto$ Electronegativity of A

ii) Strength of H-Bond $\propto \dfrac{1}{\text{Electronegativity of B}}$

Eg. Which bond has maximum hydrogen Bond strength

i) O — H - - - - - F

ii) O — H - - - - O

iii) F — H - - - F

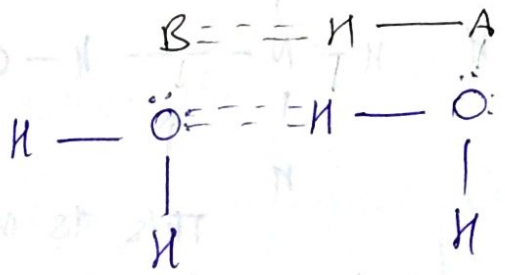
iv) F — H - - - - N

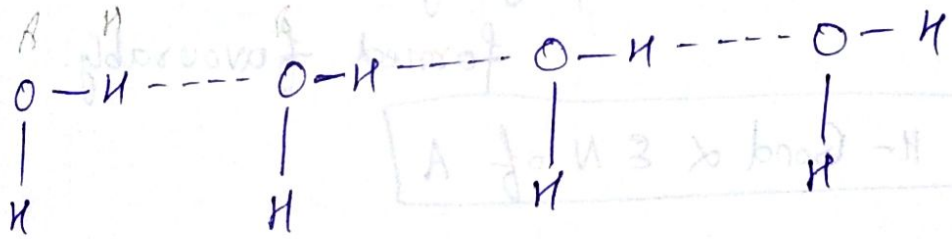
Types of Hydrogen Bonding

- **Intermolecular H-Bond** → Hydrogen bond between two different molecules
- **Intramolecular H-Bond** → Hydrogen Bond in same molecule.

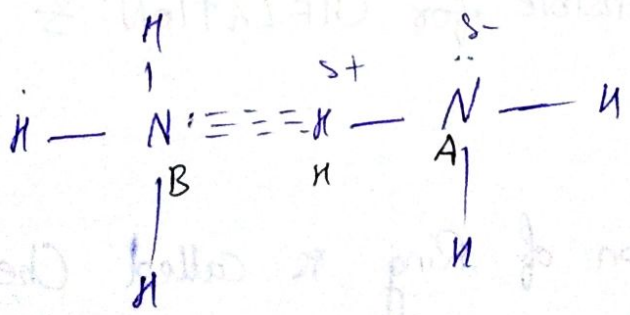
Intermolecular H-Bond

eg. H_2O & H_2O

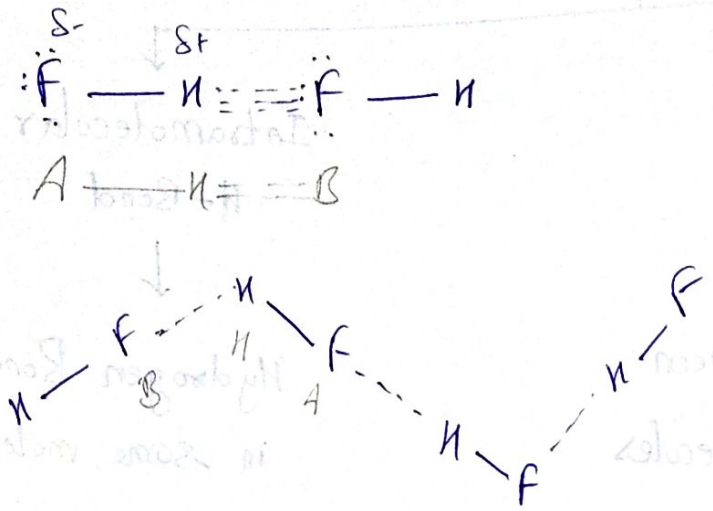




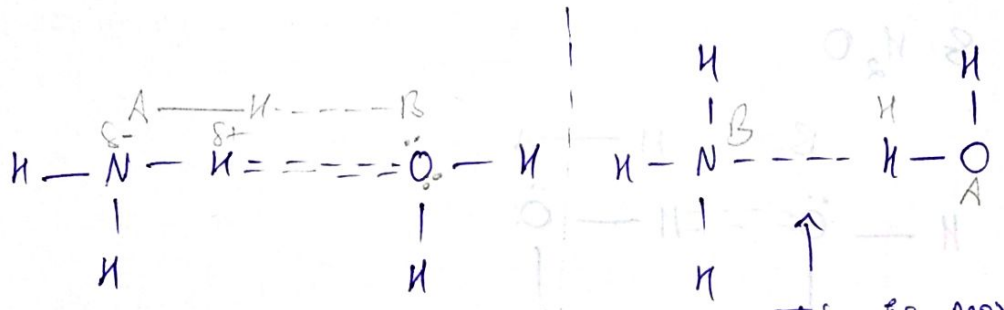
eg. NH_3 & NH_3



eg. HF & HF



eg. NH_3 & H_2O



This is more strong Hydrogen B. So, this will be formed favourably.

$\boxed{\text{• Strength of H- Bond} \propto \text{E.N of A}}$

Side Effects of H. Bond

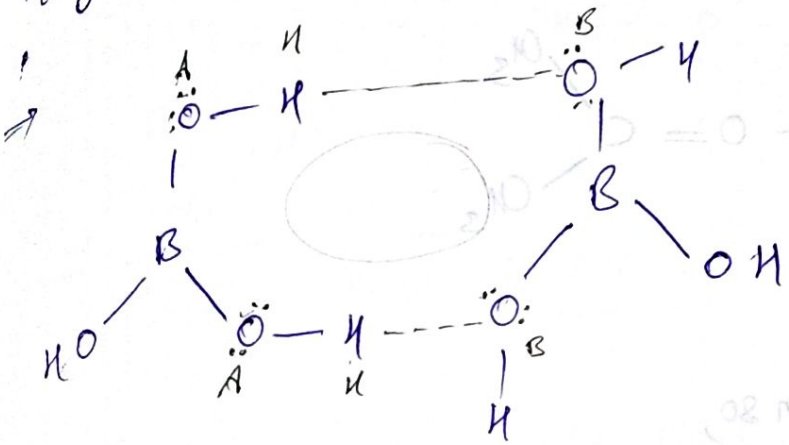
- H. Bond is also responsible for CHELATION & DIMERISATION

Chelation : Formation of Ring is called Chelation.

Dimerization: when two molecules combine to gain stability then it is called Dimerization.

1. why H_3BO_3 exists in dimer form.

$(H_3BO_3)_2$



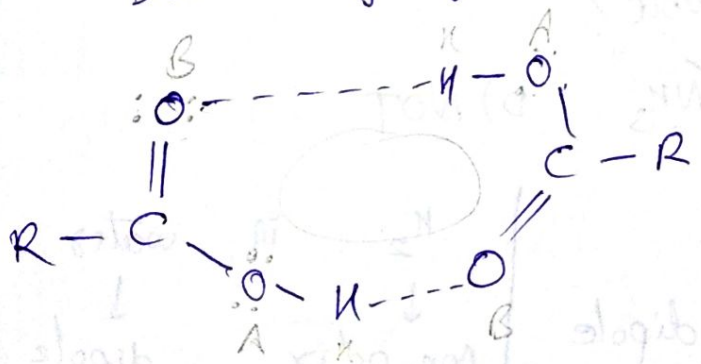
e⁻ deficient wale dimerize kar the hai!

Due to hydrogen bond, H_3BO_3 exist in dimmer form

2. Why carboxylic acid (RCOOH) exists in dimmer form

→ Due to Hydrogen Bond.

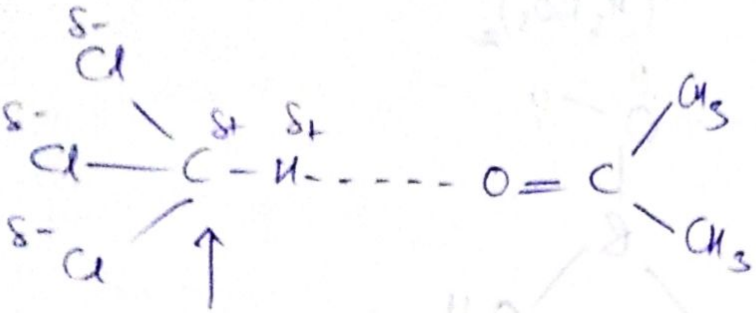
strong H bond



Special Hydrogen Bond

Choloroform & Acetone

$CHCl_3$ $\qquad$ $CH_3-\overset{O}{\overset{\|}{C}}-CH_3$

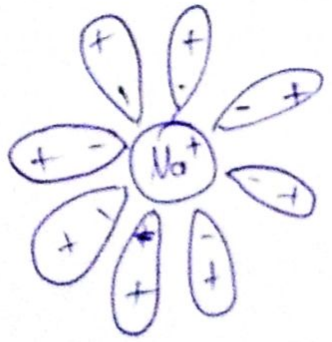


This became electronegative atom so, it attracts electrons of hydrogen towards itself

SOLUBILITY & HYDROGEN BONDING

Eg. which is more soluble?

A) Na^+ B) H_2 C) NH_3 D) NOT



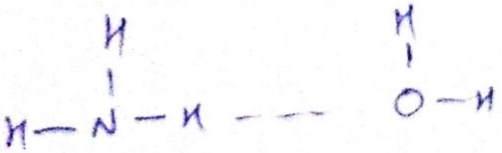
→ ion-dipole interaction

H_2 in water

↓ ↓

non-polar dipole

dipole - induced dipole

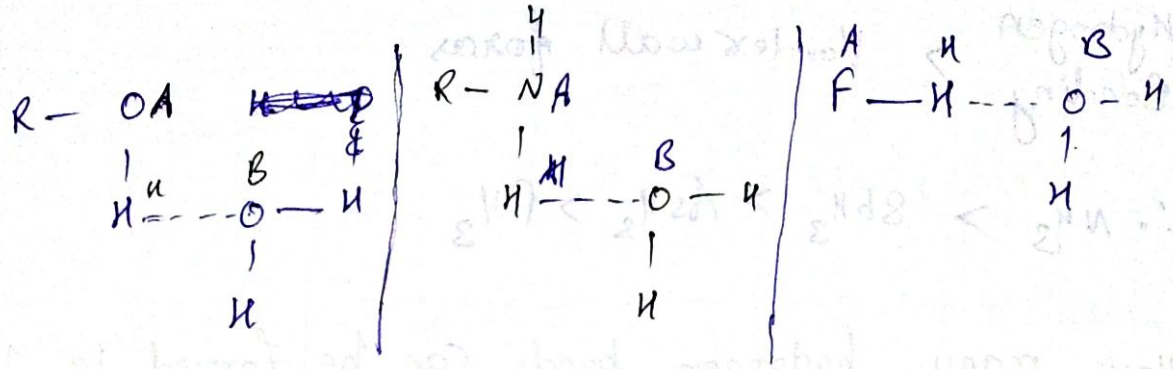


Hydrogen Bond → ∴ NH_3

Hydrogen Bond > Vander wall forces

Eg which is more soluble in water ?

A) R-OH B) R-NH_2 C) H-F D) NOT

→ 

$\boxed{\text{Stregth of H-Bond} \propto \text{E.N of A}}$

Eg. Which has more Boiling Point :

i) H_2O H_2S H_2Se H_2Te

→ H. Bonding > Vander wall force

∴ $H_2O > H_2Te > H_2Se > H_2S$

ii) HF HCl HBr HI

⇒ H. Bonding Vander wall force

HF > HI > HBr > HCl

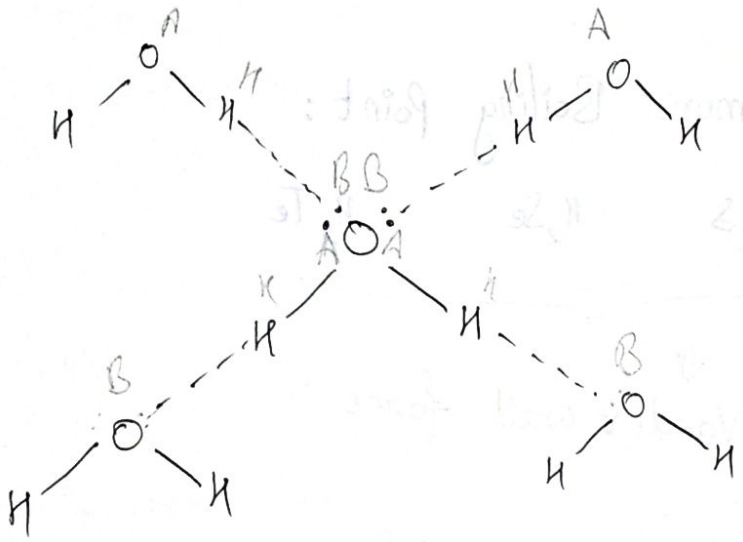
iii) $\underbrace{NH_3}_{\Downarrow}$ $\underbrace{PH_3 \quad AsH_3 \quad SbH_3}_{\Downarrow}$

Hydrogen Bonding > Vander wall forces

$\therefore NH_3 > SbH_3 > AsH_3 > PH_3$

Q) How many hydrogen bonds can be formed by 1 molecule of H_2O.

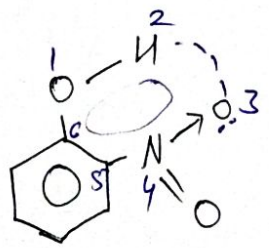
⇒



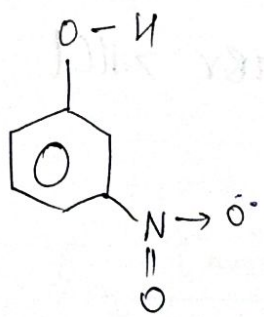
Ans: $\boxed{\text{4 - Hydrogen Bond}}$

INtra-molecule H.B (In same molecule)

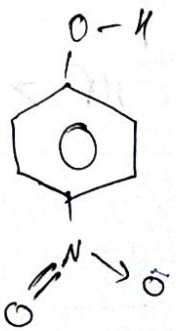
ortho nitro phenol



meta nitro phenol



Para nitro phenol



Normally, 5 to 6 membered rings are stable

Here, 6-membered ring is formed so stable hai	Here, 7-membered ring is formed so unstable hai	Here - 8 membered ring is formed so unstable hai
∴ Intra-molecule hydrogen Bonding occurs.	∴ Inter-molecule hydrogen Bonding occurs	∴ Inter molecule hydrogen Bonding occurs.

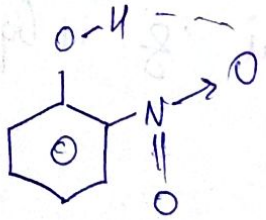
Ortho → Intra-molecular H-Bond.

Para/meta → Inter-molecular H-Bond.

Q) which has lowest Boiling point.

1) Ortho-nitrophenol
2) Meta-nitrophenol
3) Para-nitrophenol

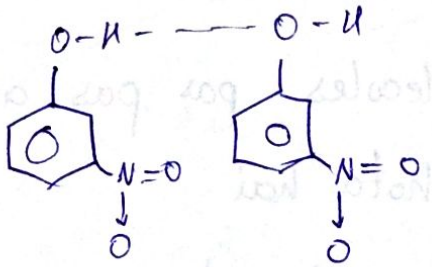
1)



intra H.B

1 molecule

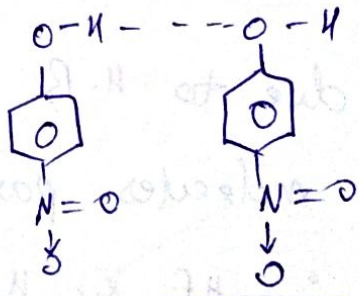
2)



inter H.B

2 molecule

3)



Inter H.B

2 molecule

B.P $\propto$ Vander wall force $\propto$ Molecular wt

∴ 1) ortho - Nitrophenol

Eg. which is more soluble in H_2O.
i) o- nitrophenol 2) m- Nitrophenol 3) P- nitrophenol

→ Compound which forms H. Bond with H_2O is soluble in it.

1)	2)	3)
Intra H.B	Inter H. Bonding	Inter H. Bonding

Eg. HF is liq while HCl is gas? H_2O is liq while H_2S is gas? Explain.

→ H-F & H_2O mai → hydrogen Bonding hota hai dono mai like ↓

due to H.B molecules pas pas atta hai & in liquid molecules pass hota hai

∴ HF & H_2O → liquid &
H_2S & HCl → gas.

Inert pair effect

Inert pair effect is applicable only in G13, G14, G15.

NOT applicable in G16, G17

As we move down in a group (13, 14, 15)

→ Lower oxidation no. ⇒ stability ↑

→ Higher oxidation no. ⇒ stability ↓

G13	G14	G15
$ns^2 np^1$	$ns^2 np^2$	$ns^2 np^3$
(+1, +3)	(+2, +4)	(+3, +5)
Boron family	Carbon family	Nitrogen family
B → +3	C → +4	N → +5
Al → +3	Si → +4	P → +5
Ga → +3 (→ ismai stable hai), +1	Ge → +4 ", +2	As → +5 ", +3
In → +3 ", +1	Sn → +4 ", +2	Sb → +5 ", +3
Tl → +3, +1 "	Pb → +4, +2 "	Bi → +5, +3 "

Why inert pair effect occurs?

$_{18}Pb \rightarrow [Xe]\ 4f^{14}\ 5d^{10}\ 6s^{2}\ 6p^{2}$

↑↓	↑	↑	

D & F subshell has very poor shielding effect.

So, 6s electrons feel more attraction of nucleus & hence, do not excite electrons.

So, in this case +2 is stable.

Inert pair effect

- The inertness of ns^2 electrons and they are not excited due to poor shielding of D & F electrons is called inert pair effect.

Eg. Which is more stable: (+3, +1)

i) $Ga^{+3} > Ga^{+1}$ ⇒ B → +3

ii) $In^{+3} > In^{+1}$ Al → +3

iii) $Tl^{+3} < Tl^{+1}$

Ga → +3 → stable hai!
Ga → +1

In → +3 → "
In → +1

Tl → +3
Tl → +1 → "

Eg which

i) Tl^{+1}

→ Tl^{+1}

Down the

ii) Tl^{+3}

→ Tl^{+3}

Q which

$PbCl_4$

→ Pb^{+4}

Unsta

↓

ab st

liya

mat

Pl

Eg Which is the stability order:

i) Tl^{+1} Ga^{+1} In^{+1} ~~Down the~~

→ $Tl^{+1} > In^{+1} > Ga^{+1}$

$\boxed{\text{Down the group} \Rightarrow \text{lower oxid}^n \text{ no.} \Rightarrow \text{stability} \uparrow}$

ii) Tl^{+3} In^{+3} Ga^{+3}

→ $Tl^{+3} < In^{+3} < Ga^{+3}$

Q Which is a better $\boxed{\text{oxidizing agent}}$

↓

दूसरों को oxidize करेगा

$PbCl_4$ or $PbCl_2$

→ Pb^{+4} Pb^{+2}

Unstable stable

↓

ab stable hone ke liya Reduce hoga matlab दूसरों को oxidize करेगा.

↓

Pb^{+2}

∴ $PbCl_4$ is better oxidizing agent