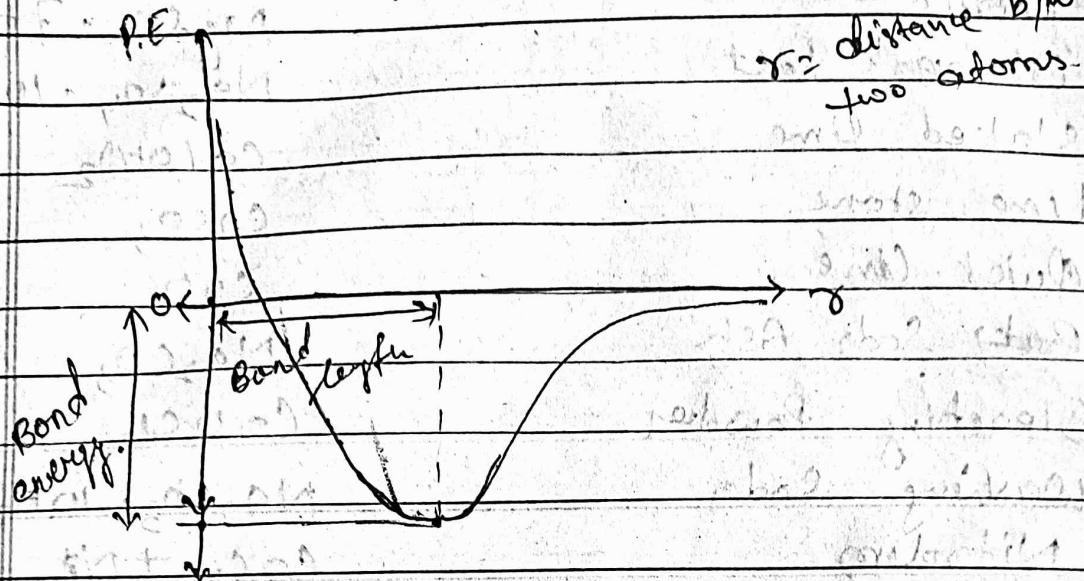


CHEMICAL BONDING① Potential Energy vs σ 

when σ becomes smaller than bond length then, nucleus-nucleus repulsion $\propto e^+ - e^-$ repulsion inc. due to which P.E. increases;

The shortest internuclear dist. b/w atoms is called bond length.

The min^m energy release during formation of bond is called bond energy.

* Chemical bond

- ① Ionic bond / Electrovalent bond
 - ② Covalent bond
 - ③ Co-ordinate bond
 - ④ Metallic bond
- * Based on B.E. $\approx 42 \text{ kJ/mol}$
- B.E. $\approx 42 \text{ kJ/mol}$

$$\rightarrow 8 \text{ kJ/mol} \leq \text{B.E} \leq 42 \text{ kJ/mol}$$

↳ Hydrogen bond

$$\rightarrow \text{B.E} \leq 8 \text{ kJ/mol}$$

weak Van der waal forces

(1) D.P - D.P (Keesom forces)

(2) D.P - ~~D.P~~^{induced} - dipole (Debye forces)

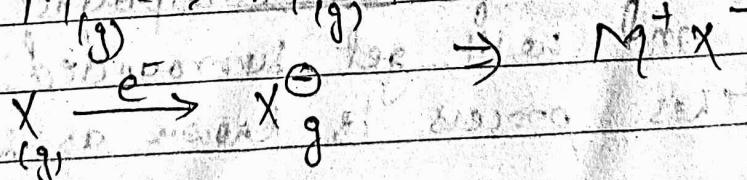
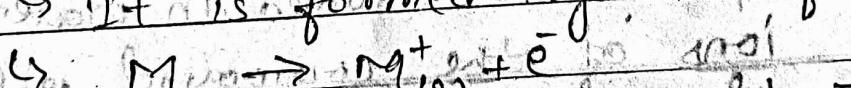
(3) T.D.P - T.D.P (London/Dispersion forces)

(4) ~~Ton~~ - T.D.P
dipole (Induced dipole)

Accn to N.C.E.R.T. ion - D.P is not considered as weak forces.

* TONIC BOND

↳ It is formed by transfer of e⁻



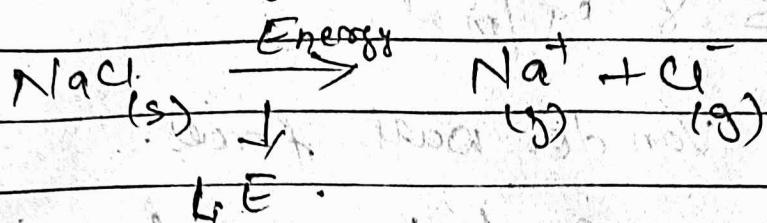
↳ favourable condⁿ for ionic bond formation.

(1) Ionisation energy of atom (metal) should be low.

(2) High ΔH_{eg} of X (non-metal)

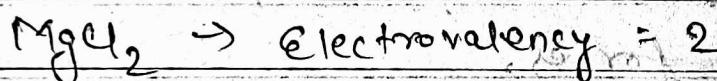
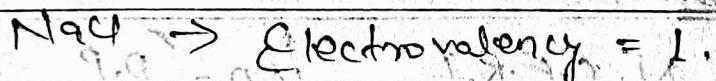
(3) The lattice energy should be high
(of M^{+}X^{-})

* The amount of energy required to break 1 mol of lattice into constituent ions in isolated gaseous state is called lattice energy.



* Electrvalency :- The no. of e⁻ transferred from M \rightarrow X.

e.g.



* Solvation or Hydration

↳ When an ionic or polar covalent compound is dissolved in water the diff. ions of the compound will get separated and will get surrounded by water molecules. This process is known as solvation or hydration.

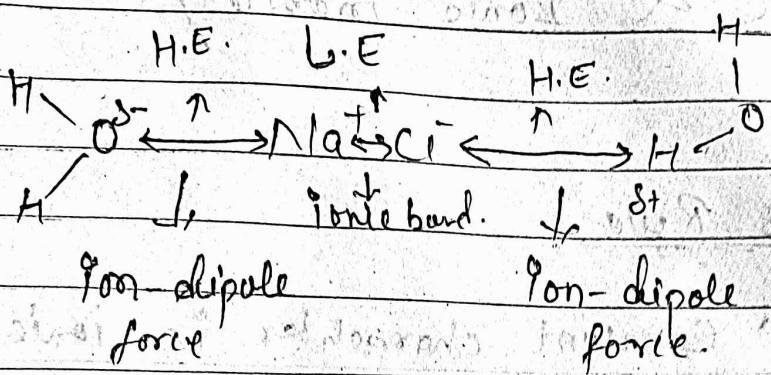
↳ The energy released during solvation is called solvation/hydration energy.

↳ The ionic compound is soluble only when H.E. $>$ L.E.

6
Date _____
Page _____

→ In solvation, Ion - dipole force acts.

e.g.



Ion-dipole
force.

* Application of solvation

① Hydrated radii

Hydrated radii $\propto \frac{1}{\text{Hydrated radii}}$

→ Solvation $\uparrow \rightarrow$ Hydrated radii \uparrow

→ Charge density $\uparrow \rightarrow$ hydration $\uparrow \rightarrow$ hydration energy \uparrow

Hydrated radii $\uparrow \leftarrow$

→ B/w size and charge, charge dominated

② Ionic mobility

→ Velocity of ions in aq. soln.

→ Solvation $\uparrow \rightarrow$ Velocity \downarrow

Ionic mobility $\propto \frac{1}{\text{Hydrated radii}}$

⑧ Electrical Conductance :-

E. C. & Ionic mobility.

* Fajan's Rule

↳ Covalent character in ionic compounds

→ Polarization :- The distortion of electron cloud of anion by the cation.

* Polarizing power :- The power of cation to attract e^- cloud of anion.

* Polarizability :- The ability of anion to get distorted by the cation.

↑ Statement :- Small cation, bigger anion and high charge on cation or anion have greater covalent character

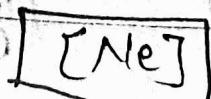
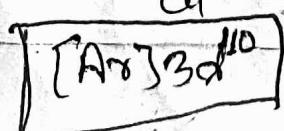
* Pseudo inert gas configuration

↳ Covalent character of Pseudo inert gas config. \rightarrow Covalent character of inert gas

e.g.

$\text{CaCl}_2 \rightarrow \text{NaCl}$ (c.c) config. (covalent character)

$\text{Ca}^+ \rightarrow \text{N}_2^+$ (P.P) \rightarrow Polarizing Power



Lecture
②

Covalent Bond

→ formed by sharing of e^- s.

(a) σ bond (Sigma)

→ $S \rightarrow S$, $S-p$, $p-p$, end-to-end or head on overlapping.

(b) π - bond

$p-p$ sidewise or lateral overlapping.

* Calculation of formal charge

→ charge on individual atoms

$$F.C. \text{ on atom} = \text{Valence } e^- - \text{lone pair } e^- - \frac{1}{2} \text{ bond. pair } e^{\pm}$$

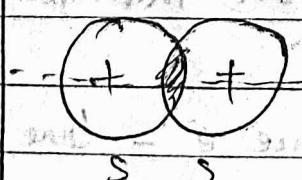
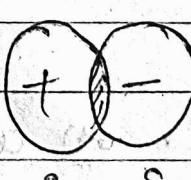
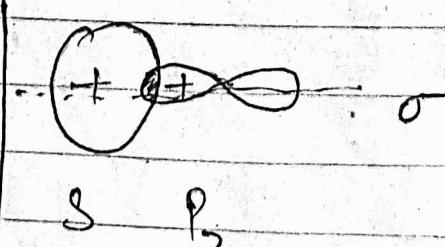
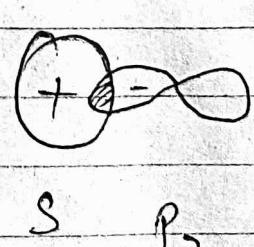
* Valence bond theory :-

Postulates :-

- ① Covalent bond is formed by overlapping of atomic orbital.
- ② Each overlapping orbital must have one unpaired electron.
- ③ When In a molecule 1st sigma bond is formed then π -bond is formed.
- ④ π -bond is weaker than sigma bonds
- ⑤ (area of overlapping = $\sigma > \pi$)
- ⑥ $A \equiv A > A = A > A - A$

Bond energy.

* Overlapping approach.

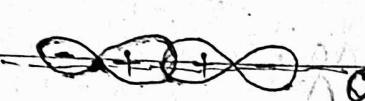
Overlapping	+ve or in phase (bond forms)	-ve or out of phase (bond doesn't form)
① S-S		
② S-P		

overlap of

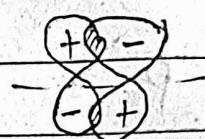
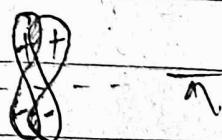
+ve

-ve

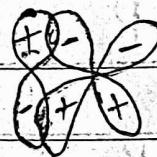
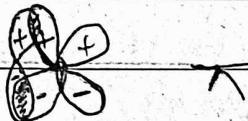
③ P-P
Head on



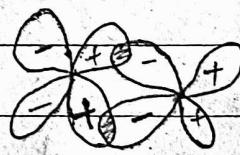
④ P-P
sidewise
or
lateral



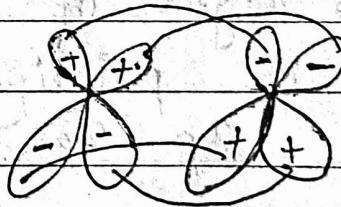
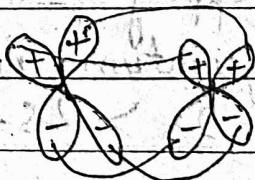
⑤ P-d
sidewise



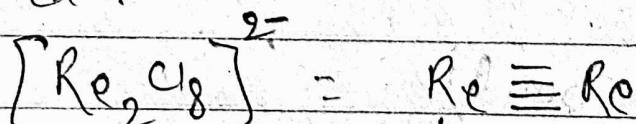
⑥ d-d
(2-lobe
interaction)



⑦ d-d
4 lobe
interaction



Ex:-



1s, 2s, 2p.

* Zero overlapping:

↳ This type of overlapping is not allowed, bond not form.

① S + P



→ nuclear distance dec.

→ repulsion inc.

② P + P

↳ $P_x - P_y$ $P_y - P_z$ $P_z - P_x$

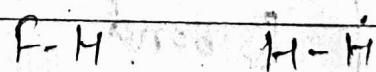
= not allowed, diff. planes.

S can not form π bond.

Strength of overlapping :- (σ bond)

If principal quantum no. same,

$$2p - 2p > 2p - 2s > 2s - 2s$$



Area of overlapping is more.

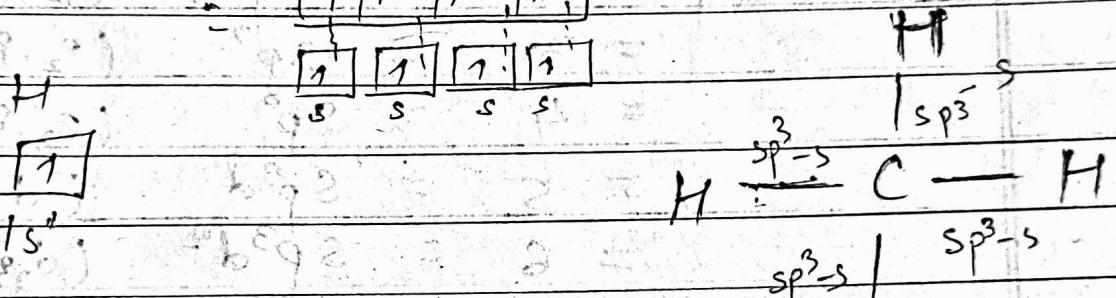
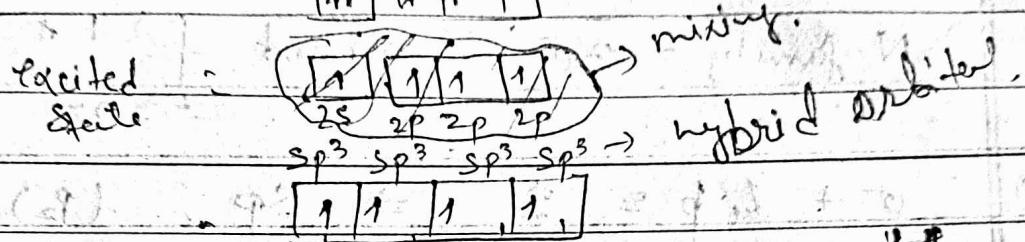
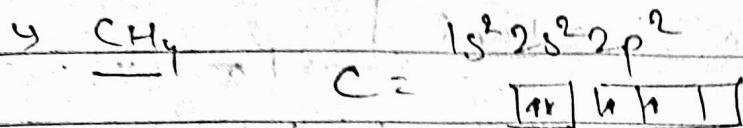
Strength (π -bond)

$$2p_{\pi} - 2p_{\pi} > 2p_{\pi} - 3d_{\pi} > 2p_{\pi} - 3p_{\pi} > 3p_{\pi} - 3p_{\pi}$$

* Hybridisation

↳ Intermixing of atomic orbital.

↳ Hybrid orbitals = Atomic orbitals mixed.

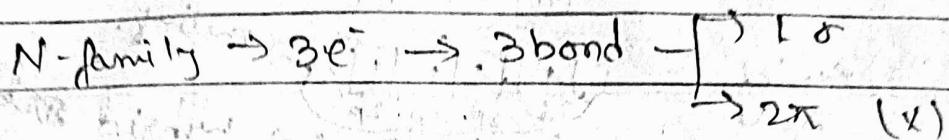
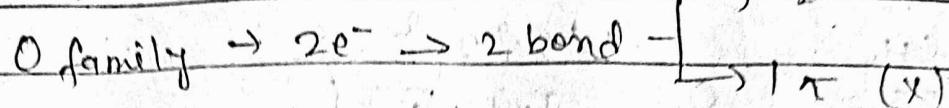
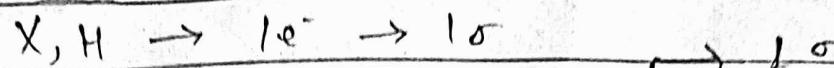


↳ Hybrid orbital =

↓ always form σ bond
can not form π -bond (generally)
(exception are there.)

To find hybridisation

- ① Find valence e⁻ of central atom
- ② Find sharing e⁻ of side atoms



- ③ $\sigma + L.p = 2 \rightarrow Sp (p_2)$
 $= 3 \rightarrow Sp^2 (p_2, p_x)$
 $= 4 \rightarrow Sp^3 (p_x, p_y, p_z)$
 $= 5 = Sp^3d (d_{z^2})$
 $= 6 = Sp^3d^2 (d_{x^2-y^2}, d_{z^2})$
 $= 7 = Sp^3d^3 (d_{x^2-y^2}, d_{z^2}, d_{xy})$

% S, P, d.

$$Sp \cdot \% S = \frac{1}{2} \times 100 = 50\%$$

$$\% P = 50\%$$

$$Sp^2 \cdot \% S = \frac{1}{3} \times 100 = 33.33\%$$

$$\% P = 66.66\%$$

$$Sp^3 \cdot \% S = \frac{1}{4} \times 100 = 25\%$$

$$\% P = 75\%$$

$$Sp^3d \cdot \% S = \frac{1}{5} \times 100 = 20\% \quad \% d = \frac{1}{5} \times 100 = 20\%$$

$$\% P = \frac{3}{5} \times 100 = 60\%$$

* Hybridisation

sp

sp^2

sp^3

sp^3d

sp^3d^2

sp^3d^3

Geometry / structure.

Exm

Linear

CO_2

Trigonal Planar

BF_3

Tetrahedral

CH_4

Trigonal bipyramidal

PCl_5

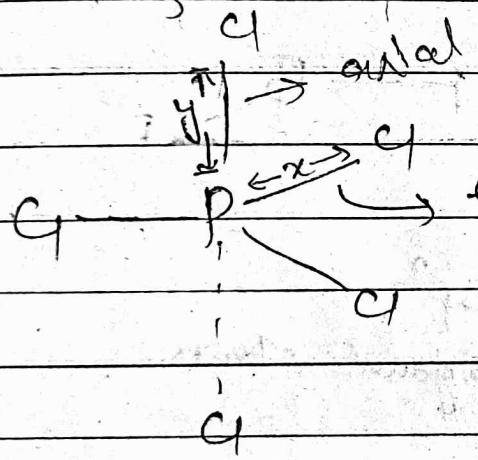
Octahedral

~~SF_6~~

Pentagonal bipyramidal. IF_7

↳ Trigonal bipyramidal.

PCl_5



$y > x$.

but in PF_5

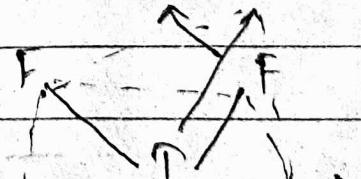
due to small size of F
peri-rotation occurs

and then $y = x$.

↳ Pentagonal bipyramidal.

IF_7

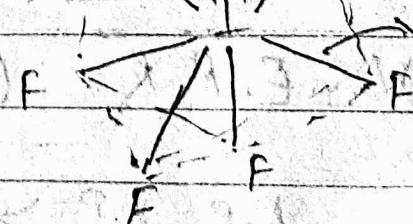
axial



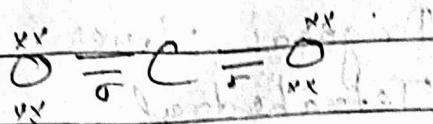
equatorial.

↳ equatorial > axial.

(orbital overlap is more greater in axial positions)



① CO_2 .



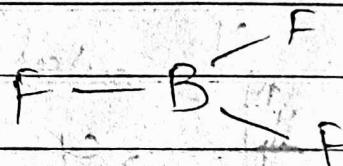
$$\sigma + \text{L.P.} = 2 + 4 \times 2 = 2 + 0 = 2.$$

(central atom) SF_6

\therefore Hybridisation = S.P.

Geometry = Linear.

② BF_3



$$\text{B} = 0 - 3$$

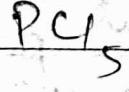
$$\text{L.P.} = 0$$

$$\therefore \sigma + \text{L.P.} = 3$$

sp^2
Trigonal planar

③

$$5 - 5 = 0$$



$$\text{L.P.} = 0$$

$$\sigma = 5$$

$$\text{sp}^3\text{d}$$

$$= \text{T.B.P}$$

* V.S.E.P.R. (Valence shell e⁻ pair Repulsion Theory)

$$\text{L.P.} - \text{L.P.} > \text{L.P.} - \text{B.P.} > \text{B.P.} - \text{B.P.}$$

\rightarrow To determine shape of compound

Compound	Hybridisation	Geometry	Shape
① XeF_2	sp^3d	T. B. P	Linear
② NH_4^+	sp^3	Tetrahedral	Tetrahedral
③ SO_4^{2-}	sp^3	Tetrahedral	Tetrahedral
④ ClF_3	sp^3d	T. B. P	T - shape
⑤ SF_4	sp^3d	T. B. P	See - saw
⑥ XeF_6	sp^3d^3	P. B. P	Distorted Octahedral
⑦ NO_2^+	sp	Linear	Linear
⑧ NO_3^-	sp^2	Trigonal planar	Trigonal planar
⑨ SO_2	sp^2	Trigonal planar	Bent

X	Y	$\text{Z} \cdot \text{P}$	Shape
2	0		Linear
2	1		Angular / Bent / V - shape
2	2		Angular / " / "
2	3		Linear
3	0		Trigonal planar
3	1		Trigonal Pyramidal
3	2		T - shape
4	0		Tetrahedral
4	1		See - saw
4	2		Square planar
5	0		T. B. P
5	1		Square Pyramidal
6	0		Octahedral
6	1		Distorted Octahedral
7	0		Pentagonal bi Pyramidal

Q. Species. Hybrid Geometry. Shape.

⑩ NO₂ Sp² Trigonal planar Linear
 $5-4=1e^-$ p lone pair Bent
 2σ
 $2\sigma + 1e^- = 3$

⑪ EOF Sp³d T.B.P T-Shape
 $8-2-2=4e^- = 2L.P.$
 3σ
 $2L.P + 3\sigma = 5$

⑫ I₃ Sp³d T.B.P Linear
 $I_1 I_2$ Linear
 $8-2=6e^- = 3L.P.$
 2σ
 $2\sigma + 3L.P = 5$

⑬ N₃⁻ sp Linear
 $N_2 N$ sp Linear Linear

$6-3\times 2=0e^- = 0L.P.$

2σ
 $2\sigma + 0L.P = 2$

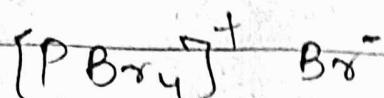
⑭ Solid PCl₅ can exist in dimer form

$\therefore [PCl_4]^+ [PCl_6]^-$ Tetrahedral Tetra-
Octahedral Octahedron

$4-4\times 2=0e^-$ sp³
 4σ
 $\therefore 4 \rightarrow sp^3$

$6-6\times 2=0e^-$
 6σ
 $\therefore 6 \rightarrow sp^3d^2$

15) Solid PBr_5
(monomer)



$$4-4=0$$

$$4\sigma = 4.$$

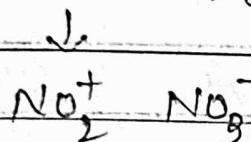
Hybrid

sp^3

Tetrahedral

Tetrahedral

16) Solid N_2O_5



$$4-4=0$$

$$2\sigma \quad 6-6=0$$

$$3\sigma$$

sp

Linear

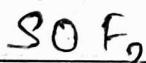
Linear

sp^2

Trigonal
planar

Trigonal
planar.

17) ~~Solid NaF~~



sp^3

Tetrahedral

Trigonal

pyramidal.

$$6-2-2=2e=1 \text{ C.P.}$$

$$3\sigma$$

$$3\sigma + 1 \text{ C.P.} = 4$$

18) ClO_3^-

sp^3

Tetrahedral

Trigonal

Pyramidal.

$$8-6=2e=1 \text{ C.P.}$$

$$3\sigma + 1 \text{ C.P.} = 4$$

19) SeO_2F_2

sp^3d

T.B.P

See-saw

$$8-4-2=2e=1 \text{ C.P.}$$

$$4\sigma$$

$$4\sigma + 1 \text{ C.P.} = 5$$

20) XeF_4

sp^3d^2

Octahedral

Square
Planar

$$8 - 4 \times 2 \times 6 = 26 \text{ P.}$$

4σ

$$4\sigma + 21 \text{ P.} = 6$$

21) $\text{H}_2\text{O} + \text{H}^+ \rightarrow \text{H}_3\text{O}^+$ sp^3 Tetrahedral Trigonal
Pyramidal

$$5 - 3 = 2e = 1 \text{ L.P.}$$

$$3\sigma + 3\sigma + 1 = 4$$

$$6 - 2 - 4 = 2 \text{ L.P.}$$

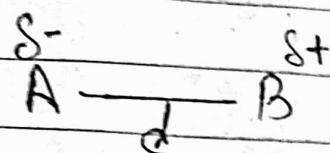
2σ 2

22) $\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$ sp^3 Tetrahedral Tetrahedral

$$4 - 4 = 0e = 0 \text{ L.P.}$$

$$4\sigma \quad 4\sigma$$

* Dipole moment (μ) :- It is the product of (+ve) or (-ve) charge and distance b/w two atoms.



d = separation b/w two charges.

$$[\mu = q \times d]$$

q = charge.

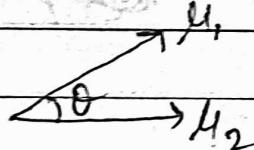
→ μ is a vector quantity.



if $E_N A > E_N B$.

$$\therefore A \leftarrow + B$$

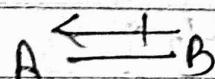
→ Resultant of vector



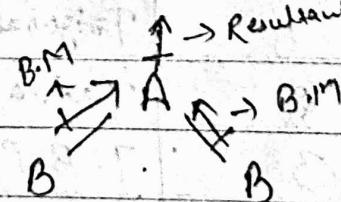
$$M_{res} = \sqrt{l_1^2 + l_2^2 + 2l_1 l_2 \cos\theta}$$

→ Bond moment :- Dipole moment of a bond is called bond moment.

e.g.



$$B.M = D.M.$$



D.M of individual bond = B.M

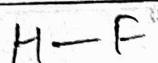
Resultant of bond moment = D.M.

↳ Unit of μ = 1 Debye (D)

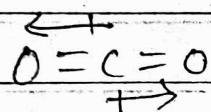
→ If μ of compound = 0; non polar molecule
If $\mu \neq 0$; polar molecule



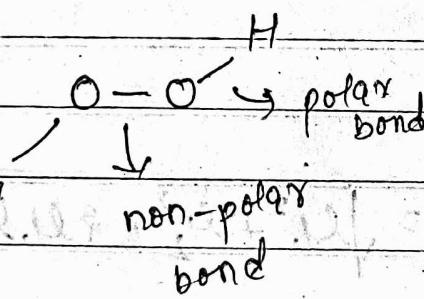
E.N. diff. = 0. non-polar bond; non polar molecule



E.N. diff. $\neq 0$ polar bond, polar molecule



polar bond; non polar molecule

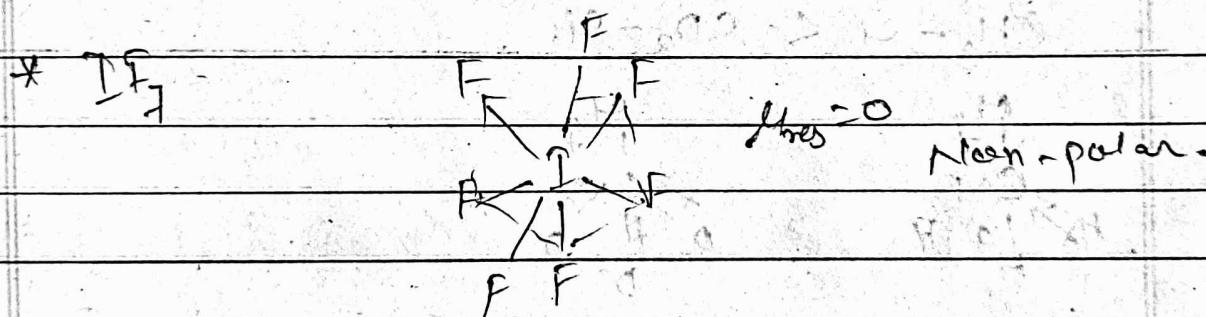
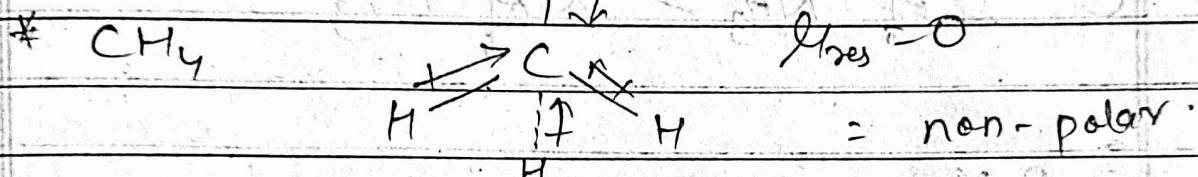


$\mu_{\text{res}} \neq 0$

∴ non-polar molecule

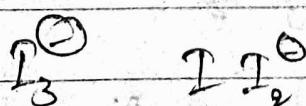
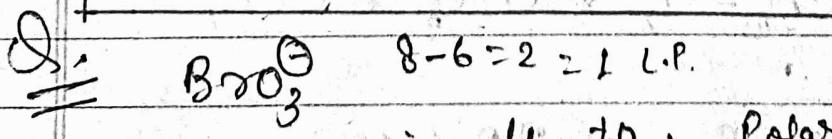
σ bond	L.P	Geometry	Example
0	0	Linear	$\text{BeCl}_2, \text{CO}_2$
2	1	Tetrahedral	SO_2
2	2	Tetrahedral	H_2O
2	3	T.B.P	XeF_2
3	0	Trigonal planar	BF_3
3	1	Tetrahedral	NH_3
3	2	T.B.P	ClF_3
4	0	Tetrahedral	CH_4
4	1	T.B.P	SF_4
4	2	Octahedral	XeF_4

5	0	T.B.P	PCl ₅
5	1	Octahedral	IF ₅
6	0	Octahedral	SF ₆
6	1	P.B.P (Distorted octahedral)	XeF ₆
7	0	P.B.P	IF ₇



* If L.P = 0 with same side atom
 $\Rightarrow \mu_{res} = 0 \Rightarrow$ N.P.

* B.P L.P with same side atom
2 3 4 2 $\Rightarrow \mu_{res} = 0 \Rightarrow$ N.P.



$$8-2=6 \neq 3 \cdot 1 \cdot 2$$

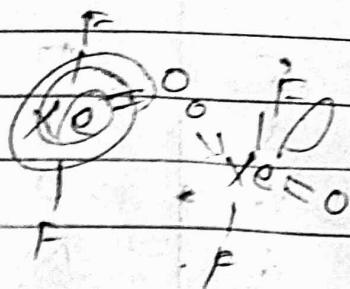
2, 3, $\mu_{res} = 0$, non-polar.

Q XeO_2F_2

$$8-4-2 = 2e^- \text{ L.P.}$$

4σ, 1. L.P. $\mu_{\text{net}} \neq 0$

Polar.

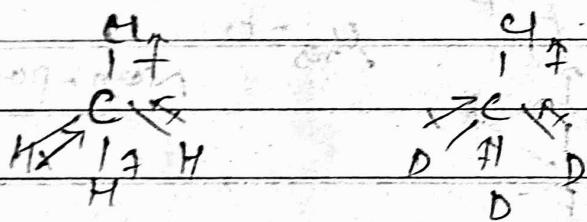


* Some important examples of H.

① $\text{H of } \text{NH}_3 > \text{H of } \text{NF}_3$

② D.M.

$\text{CH}_3-\text{Cl} < \text{CD}_3-\text{Cl}$



$+2 \text{ of } \text{D} > +2 \text{ of } \text{H}$

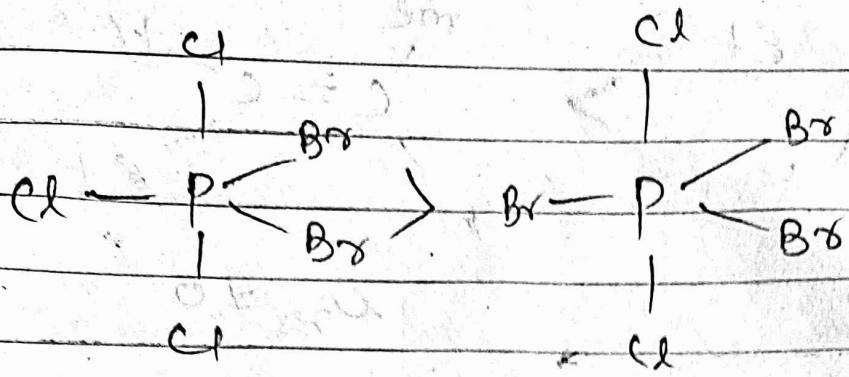
③ $\text{CH}_3-\text{Cl} > \text{CH}_3-\text{F} > \text{CH}_3-\text{Br} > \text{CH}_3-\text{I}$

$\therefore \text{H of } (\text{C-Cl}) > \text{H of } (\text{C-F})$

④ $\text{CH}_3-\text{Cl} > \text{CH}_2-\text{Cl}_2 > \text{CHCl}_3 > \text{CCl}_4$

⑤ D.M. $\text{PCl}_3\text{Br}_2 > \text{PBr}_3\text{Cl}_2$

* Bent rule:- more electronegative atoms must be present on axial position in T.B.P geometry



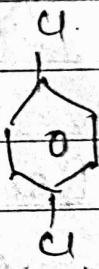
$\mu_{\text{res}} \neq 0$
polar

$\mu_{\text{res}} = 0$
non-polar

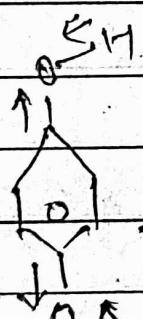
⑥ D.M. of



$\neq 0$ but D.M. of

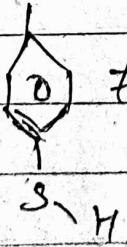


$= 0$



$\mu_{\text{res}} \neq 0$

⑦ D.M. of

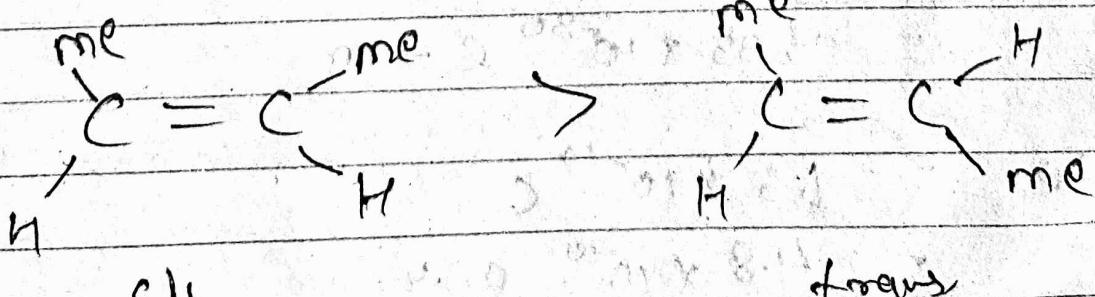


but D.M. of



$= 0$

⑧

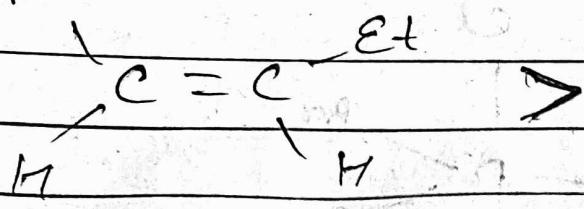


$\mu_{\text{res}} \neq 0$

forwards

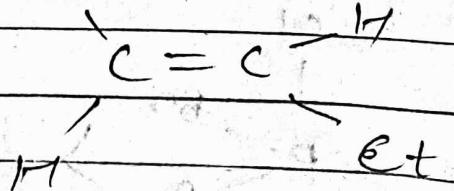
$\mu_{\text{res}} = 0$

me

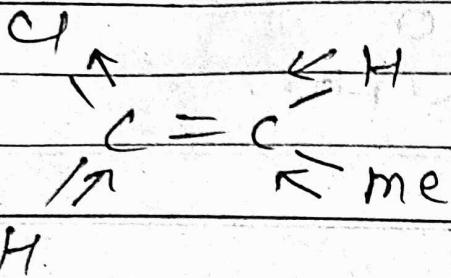


$\mu_{res} \neq 0$

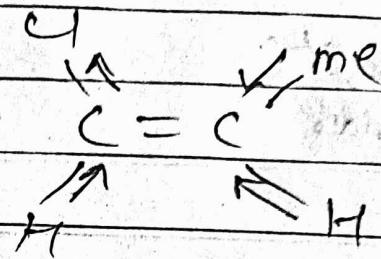
me



$\mu_{res} \neq 0$



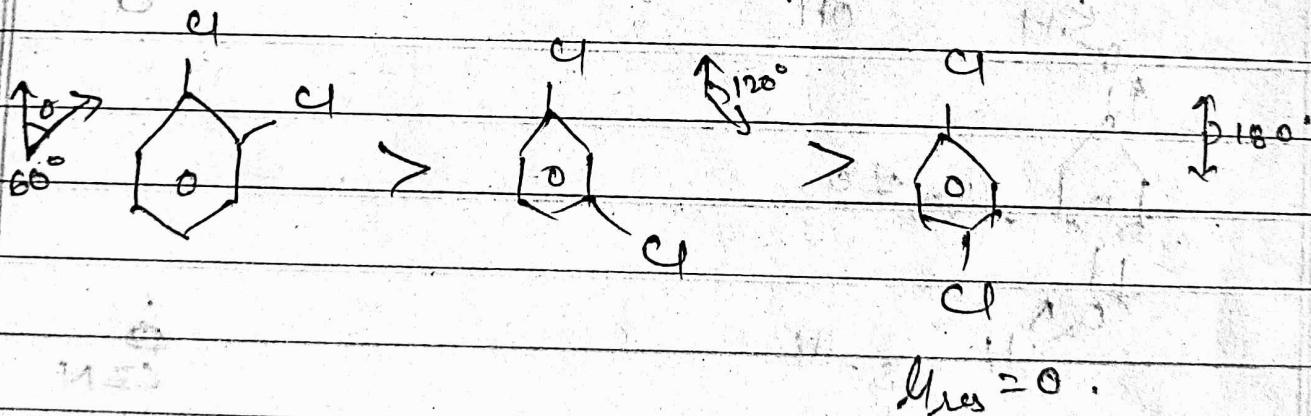
trans



cis

⑨

Benzene Derivatives



$\mu_{res} = 0$

* Unit conversion.

$$1 \text{ D} = 10^{-18} \text{ esu-cm}$$

$$= 3.33 \times 10^{-30} \text{ C-m}$$

$$1 \text{ e} = 1.6 \times 10^{-19} \text{ C}$$

$$= 4.8 \times 10^{-10} \text{ e.s.u}$$

* Calculation of % ionic character.

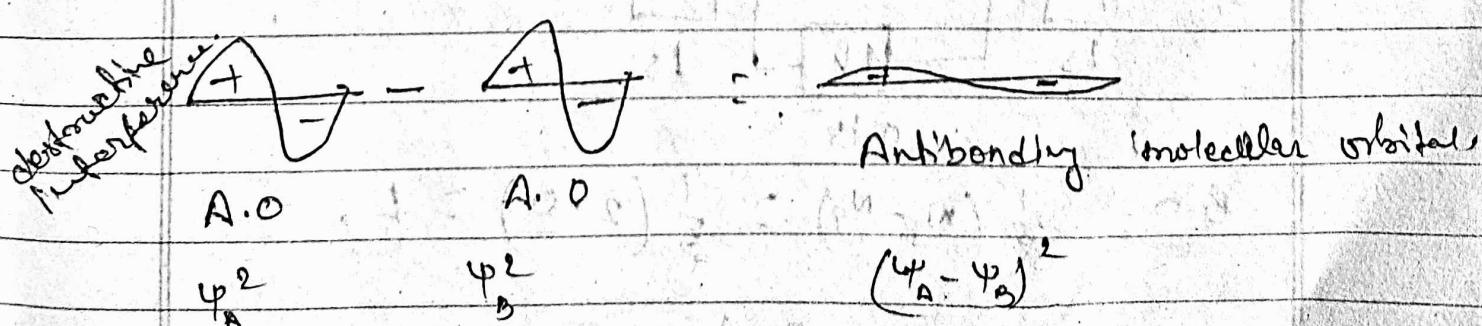
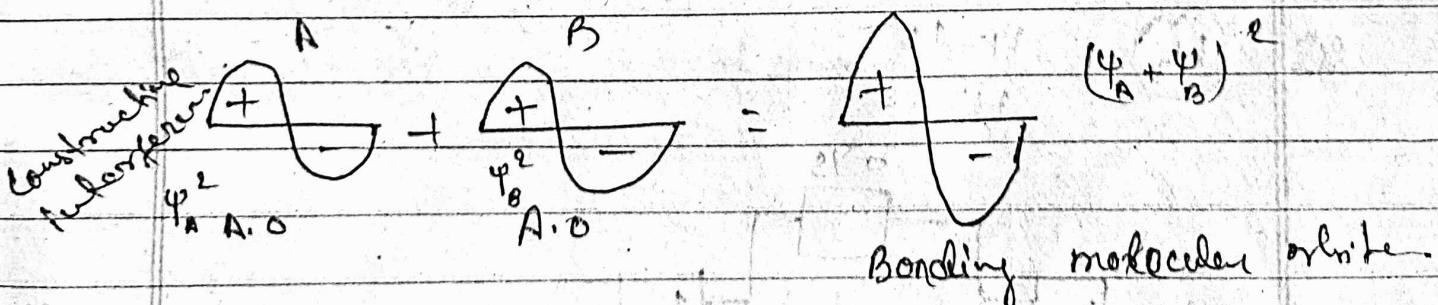
$$\% \text{ I.C.} = \frac{\text{Electrons}}{\text{Mol.}} \times 100$$

Factor 5

Molecular Orbital Theory.

* Postulates

- ① In V.B.T. only valence e⁻s will take part in bond formation but in M.O.T. all e⁻s will take part in bond formation.
- ② Molecular orbitals are formed by combination of atomic orbitals.
- ③ L.C.A.O. method (Linear combination of atomic orbitals)



(4) Fillings of σ & π in M.O. is same as that of Atomic Orbitals.
(Follow Aufbau's, Hund's and Pauli's)

(5) Bonding molecular orbital has lower energy, (more stable)

(6) Anti-bonding M.O. has higher energy. (less stable)

(7) B.O (Bond order)

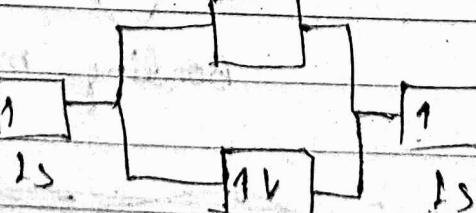
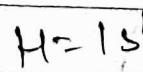
$$\text{B.O} = \frac{1}{2} (N_b - N_a)$$

B.O $\uparrow \rightarrow$ B.E $\uparrow \rightarrow$ B.L \downarrow

B.O $\uparrow \rightarrow$ Stability \uparrow

(8) magnetic nature: Unpaired e's \rightarrow paramagnetic
paired e's \rightarrow diamagnetic

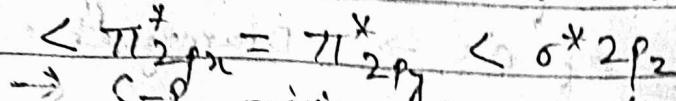
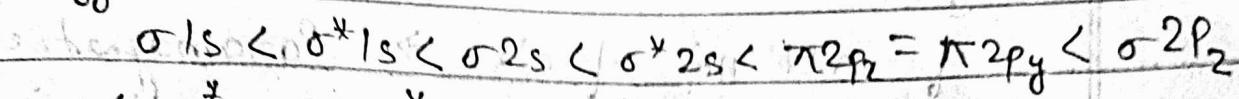
e.g.



$$\text{B.O} = \frac{1}{2} (N_b - N_a) = \frac{1}{2} (2 - 0) = 1.$$

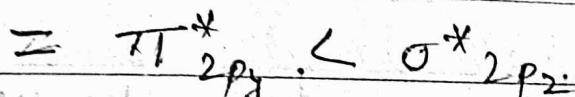
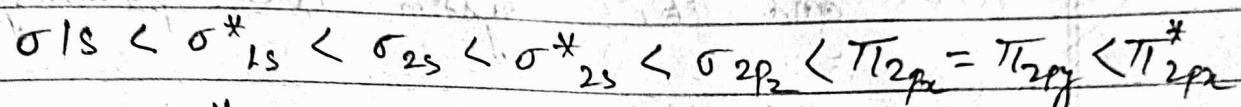
$$\text{B.O} = 1.$$

Energy Order :- Case ① (when no. of e^- 's ≤ 14)



\rightarrow S-p mixing occurs ($\because \pi 2p_z, \pi^* 2p_y$ comes before $\sigma^* 2p_z$)

case ② (when no. of e^- 's > 14)



\rightarrow S-p mixing does not occurs ($\because \sigma^* 2p_z$ comes before $\pi^* 2p_z, \pi 2p_z$).

* Super trick.

$y \rightarrow$ bond order

\rightarrow case ① if no. of e^- 's ≥ 8

$x \rightarrow$ no. of e^- 's

0	0.5	1	1.5	2	2.5
8	9	10	11	12	13

P, S P

D P

Aug molecule

14 B

2.5	2	1.5	1	0.5	0
15	16	17	18	19	20

\rightarrow case ② if no. of e^- 's ≤ 8 .

B A B A

ex^m.

$$Li_2 = 3e^- + 3e^- = 6e^-$$

2 2 2

B A B A

$$B.O. = \frac{1}{2} (N_b - N_a) = \frac{1}{2} (4 - 2) = 1.$$

* Magnetic nature.

if odd no. of e's \rightarrow Paramagnetic.

if even no. of e's \rightarrow Diamagnetic.

except 10e's and 16e's systems.

10e's and 16e's systems.

10e's and 16e's system are paramagnetic.

Question

Species	No. of e's	B.O.	magnetic nature.
B_2	10	1	Para
B_2^+	9	0.5	Para
B_2^-	11	1.5	Para
C_2	12	2.2 ($C=C$) Both π	Dia
C_2^+	11	1.5	Para
C_2^-	13	2.5	Para
N_2	14	3	Dia
N_2^+	13	2.5	Para
N_2^-	15	2.5	Para
BC	11	1.5	Para
BC^+	10	1	Para
BC^-	12	2	Dia
O_2	16	2	P
O_2^+	15	2.5	P
O_2^-	17	1.5	P
O_2^{2+}	14	3	D
O_2^{2-}	18	4	D
F_2	18	1	D
F_2^+	17	1.5	P
F_2^-	19	0.5	P

CO	14	3	D
co CO⁺	13	(2.5) 3.5	D
CN ⁻	14	3	D

Q. Compare the stability.

	Li ₂	Li ₂ ⁺	Li ₂ ⁻
No. of e ⁻	6	5	7
B.O.	2.2	2.1	2.2

$$B.O. = \frac{1}{2}(4-2) = 1 \quad \frac{1}{2}(3-2) = 0.5 \quad \frac{1}{2}(4-3) = 0.5$$

Stability \propto B.O.

↳ If B.O. is same then least e⁻ imp. B.M.O. is more stable.

$$\therefore Li_2 > Li_2^+ > Li_2^- \quad (\text{stability}).$$

~~stability~~

Q. Compare bond length.

	O ₂	O ₂ ⁺	O ₂ ⁻	O ₂ ²⁻
No. of e ⁻	16	15	17	18
B.O.	2	2.5	1.5	1

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

$$\therefore O_2^+ < O_2 < O_2^- < O_2^{2-} \quad (\text{bond length})$$

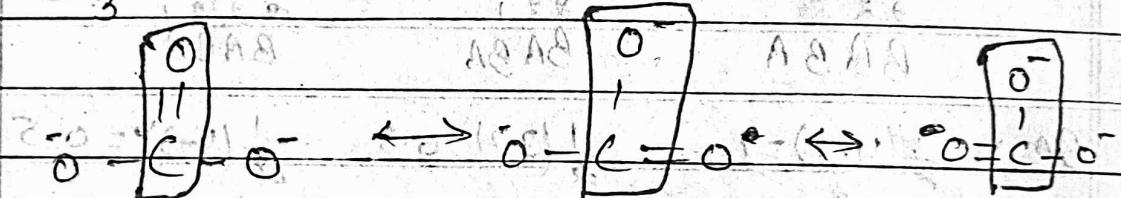
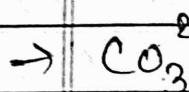
* Special case.

If no. of e's ≥ 20 .

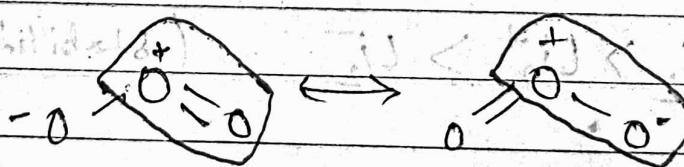
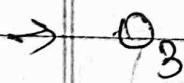
↳ by Resonance.

B.O = No. of bonds b/w two atoms in all the structures

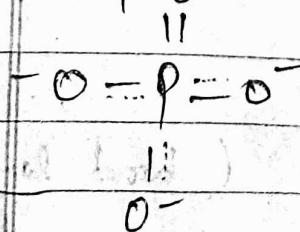
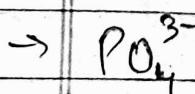
↳ no. of Resonating structure.



$$\text{B.O} = \frac{2+1+1}{3} = \frac{4}{3} \approx 1.33$$

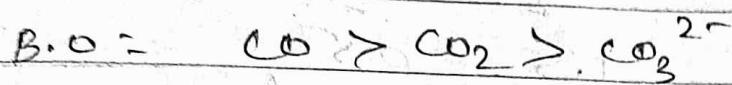


$$\text{B.O} = \frac{2+1}{2} = \frac{3}{2} = 1.5$$



$$\text{B.O} = \frac{2+1+1+1}{4} = \frac{5}{4} = 1.25$$

$$\begin{array}{cccc}
 \text{Q:} & \text{CO} & \text{CO}_2 & \text{CO}_3^{2-} \\
 \text{no. of e}^- & 14 & 1 & \\
 \text{B.o.} & 3 & & 6 - C = 0 \\
 & & & C = 6 - 3 \\
 & & & = 3 \\
 & & & 2 \\
 & & & 11 \\
 & & & - 0 \\
 & & & \hline
 & & & 11 \\
 & & & \hline
 & & & 3 \\
 & & & \frac{11}{3} \\
 & & & 3.67
 \end{array}$$



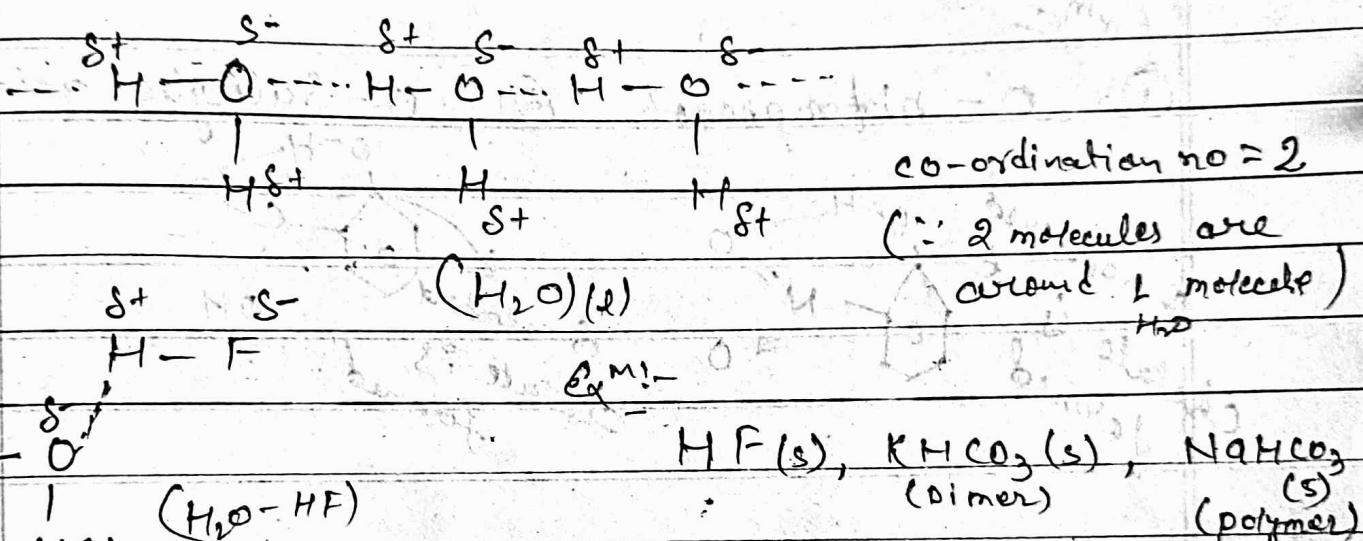
~~Feature~~ (6) * Hydrogen bonding.

- ↳ The Hydrogen bond is shown by those covalent compounds in which Hydrogen is directly bonded with more electronegative elements like F, O, N, etc.
- ↳ In some cases C and Cl also involved in H-bond.

↳ It is weak force. ($8-42 \text{ kg/mol}$ energy released)

Types.

① Intermolecular H-bonding:
Whenever H-bonding occurs b/w two
same or different species then H-bond is
called intermolecular H-bond.



~~SH~~ (A) --- H --- (B)

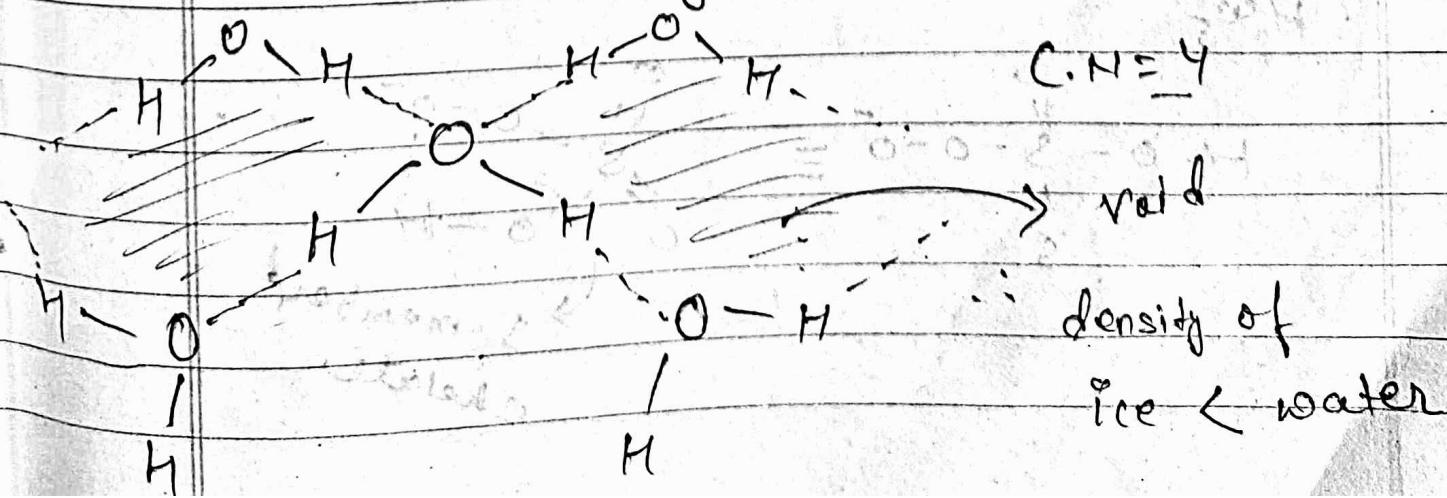
$A, B \in [F, O, N]$ in H-bonding.

lie.

$\text{O} \cdots \text{H} - \text{S}$ \neq H-bonding

$$O-H-N = \text{H-bonding.}$$

* H_2O (s) = Cage like structure.



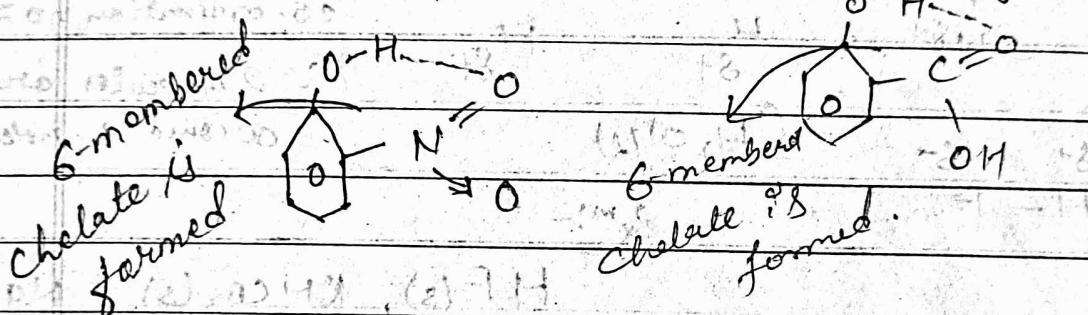
* Intramolecular H-bonding:-

↳ Whenever H-bond occurs ~~b/w~~ within a molecule then H-bond is called intramolecular H-bonding.

↳ Chelation occurs (formation of ring).

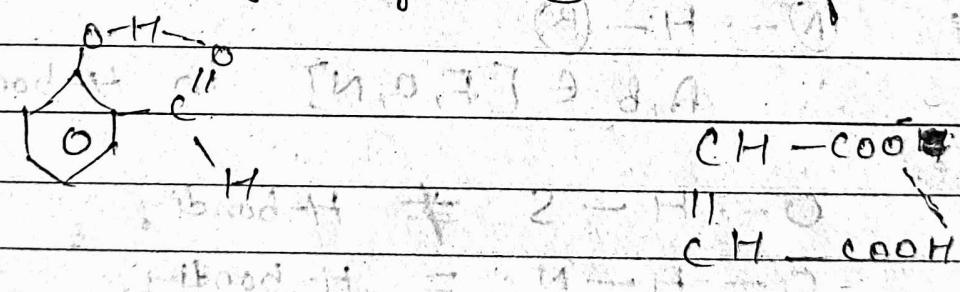
Exm:-

① O-nitrophenol ② O-Salicylic acid.

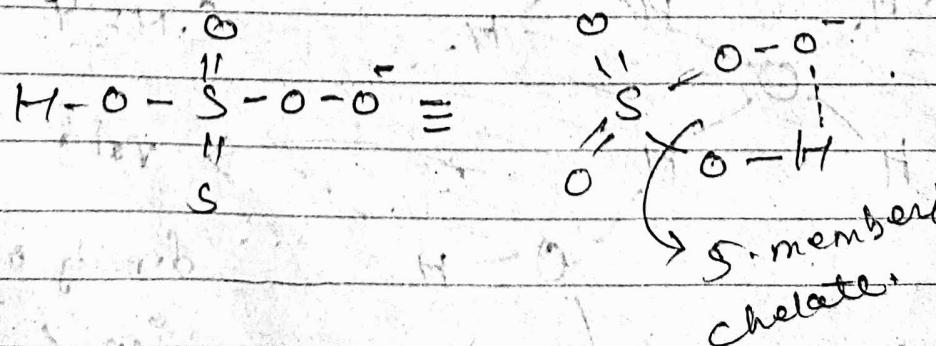
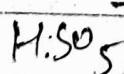


③ O-Salicyraldehyde

④ Anion of Maleic acid



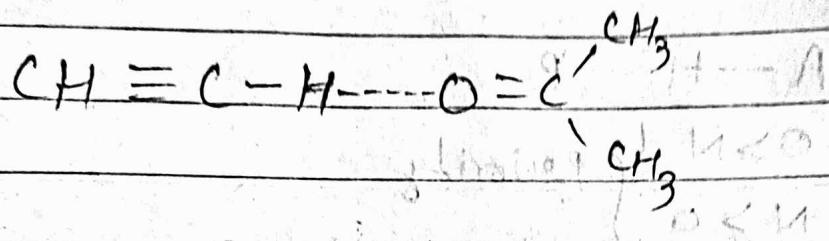
⑤ Anion of Caro's acid.



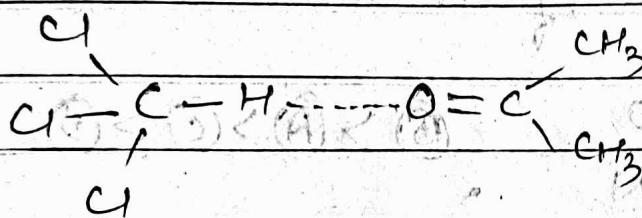
* Special cases

① Carbon

ⓐ Acetylene and acetone

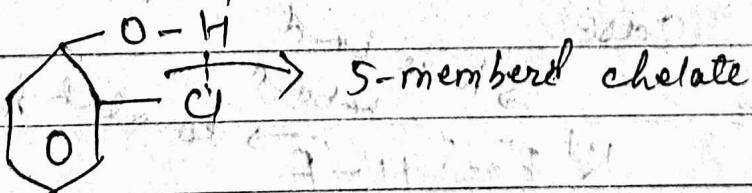


ⓑ Chloroform and acetone

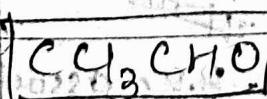


ⓐ Chlorine.

① O-chlorophenol

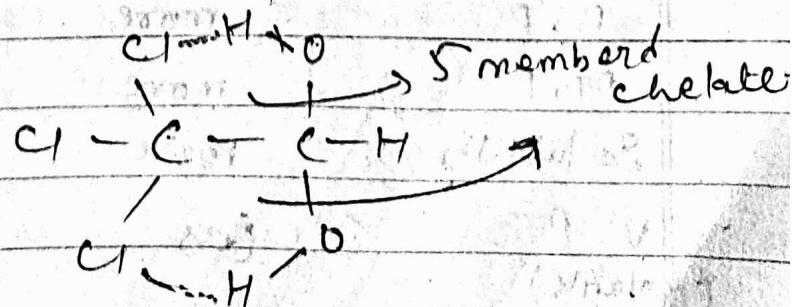


② chloral (Hydrated form)



Hydrated form =

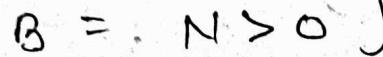
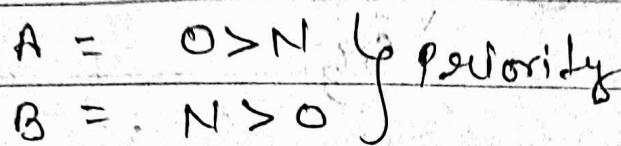
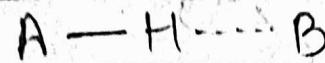
Chloral



* Strength of H-bonding.

① If F is involved, bond energy is maxm.

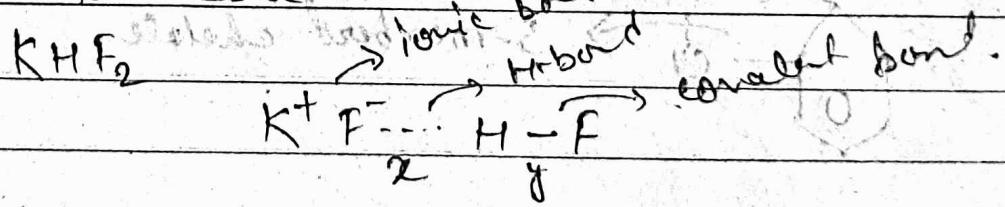
② If O and N are involved, then



Q. Compare strength of H-bond

- A) O-H---O B) O-H---N C) N-H---N D) N-H---O
- (B) > (A) > (C) > (D)

* Special case



bond energy $x = y$

* Application of H-bonding.

Properties

Inter

Intra

B.P of association

B.P

more

less

M.P & C.L.E

M.P

more

less

Solubility of association

Solubility

more

less

U.P & T

V.P

less

more

but

(relative)

less

more

N.P & B.P

* Clathrates \rightarrow inert gas trapped in cage or voids of ice.

∴ ice is capable of forming clathrates.

He and Ne do not form clathrates due to very small size, while noble gas followed by Ar forms clathrates.

Basicity of Amines in aq. medium.

Basicity, $2^\circ > 1^\circ > 3^\circ$ when $R = \text{me}$ $1^\circ = R-\text{NH}_2$

Basicity, $2^\circ > 3^\circ > 1^\circ$ when $R = \text{Et}$ $2^\circ = R_2-\text{NH}$ $3^\circ = R_3-\text{N}$

where $R \notin \text{ame, Et}$.

Then

Basicity, $3^\circ > 2^\circ > 1^\circ$

* Bent's Rule: - (sp³d hybridisation)

↳ The lone pair in T.B.P geometry must be placed on equatorial position and one e.n. element on axial position.

* Drago's Rule:-

↳ When central atom belongs to 3rd period and beyond and E.N of side atom is 2.1 or ~~more~~ less. (H).

↳ Central atom does not take part in hybridisation.

Ex^m:

$\text{PH}_3 \rightarrow \text{No hybridisation}$

Hydrides of
0 15th group

bond angle = NH_3 , $\boxed{\text{PH}_3 \text{ AsH}_3 \text{ SbH}_3 \text{ BiH}_3}$

Neatly 90°
u top to bottom
B.A dec.

107° sp^3 No hybridisation

u 16th group

bond angle = OH_2 $\boxed{\text{H}_2\text{S} \text{ H}_2\text{Se} \text{ H}_2\text{Te} \text{ H}_2\text{Po}}$

No hybridisation

angle 104.5° sp^3

near 90°

u top to bottom B.A dec.

B.A of 15th group > B.A of 16th group.

* How to compare bond angle.

① Hybridisation of central atom.

$\text{sp} > \text{sp}^2 > \text{sp}^3$ i.e. $\text{CO}_2 > \text{BF}_3 > \text{CH}_4$

(-318°, -120°, 109.5°)

Linear Trigonal Tetrahedral

② L.P. on central atom

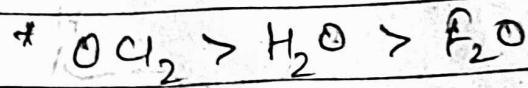
B.A $\propto \frac{1}{1 \cdot P}$ i.e. $\text{CH}_3 > \text{NH}_3 > \text{H}_2\text{O}$
 $\rightarrow \text{sp}^3 \text{ sp}^3 \text{ sp}^3$
 $1 \cdot P = 0 \quad 1 \quad 2$

③ E.N. of central atom

$\boxed{\text{B.A} \propto \text{E.N.}}$

IV. $\text{NH}_3 > \text{PH}_3 > \text{AsH}_3 > \text{SbH}_3 > \text{BiH}_3$

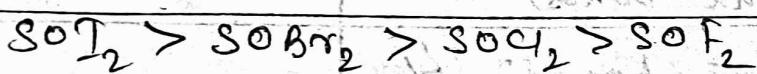
④ Size of side atom $\text{OCl}_2 > \text{OF}_2$
(B.A & size)



Q. Compare B.A.

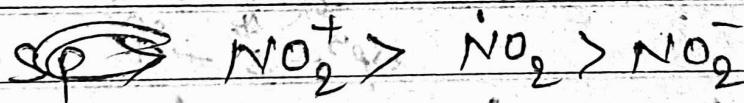
	SOF_2	SOCl_2	SOBr_2	SOT_2
hybrid.	sp^3	sp^3	sp^3	sp^3
L.P.	$1.2 < 1.4 < 1.8$	1.2	1.2	1.2

Size of side atom



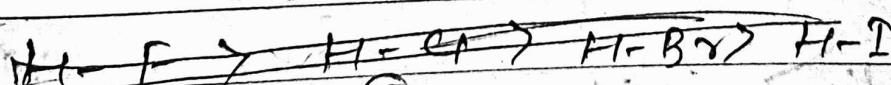
⑤ $\text{NO}_2^+ \text{NO}_2 \text{NO}_2^-$

Hybr.: sp^2 sp^2 sp^2 \hookrightarrow more repulsion in L.P. \rightarrow $\text{NO}_2^+ > \text{NO}_2 > \text{NO}_2^-$

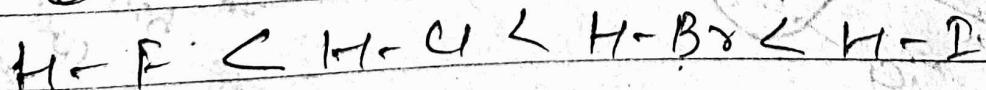


* How to compare bond length

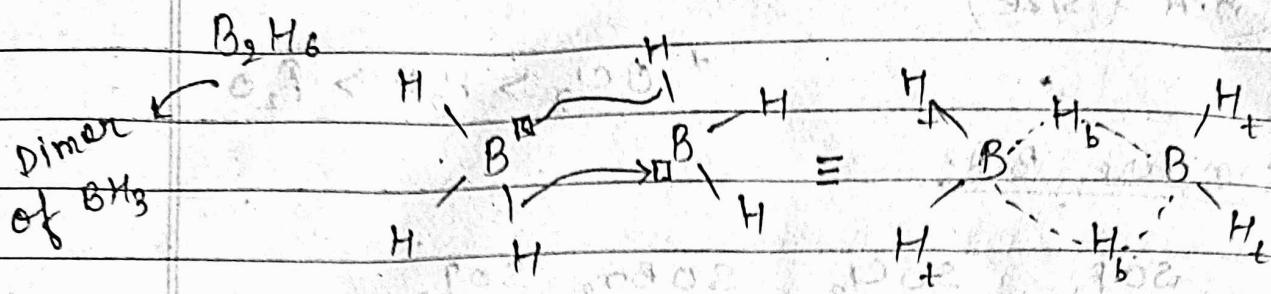
→ Size diff \uparrow , Extent of overlapping \downarrow , B.L. \uparrow



⑥ *

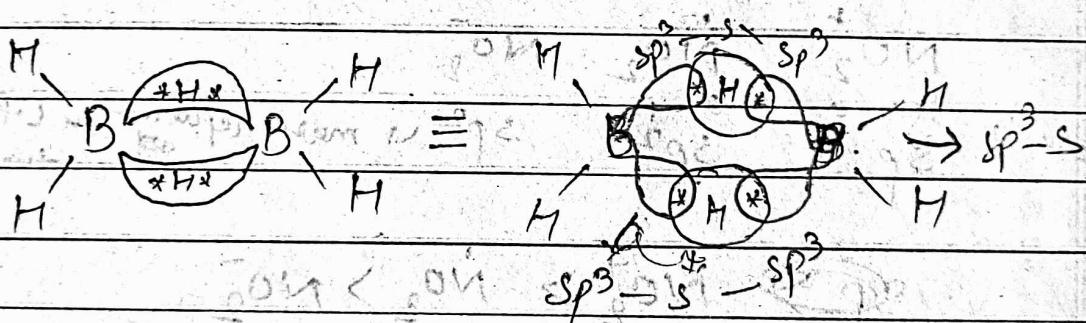


* Electron deficient bond Banana bond



↳ Hybridisation of B in $BH_3 \rightarrow sp^2$
 " " B " $B_2H_6 \rightarrow sp^3$

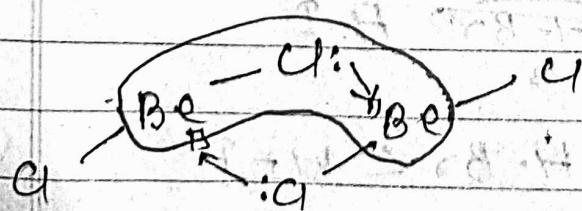
↳ $3C-2e^-$ bonds = 2 (banana bonds) \rightarrow banana bond length \rightarrow covalent bond
 ↳ $3C-2e^-$ bond = 4



Ex:- $Al_2(CH_3)_6$, $(BeH_2)_n$, $(MgH_2)_n$

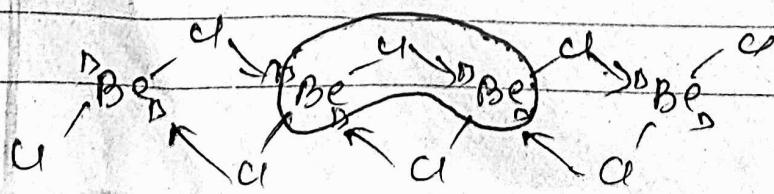
* $3C-4e^-$ bond

$\rightarrow BeCl_2$ Vapour phase (dimer)



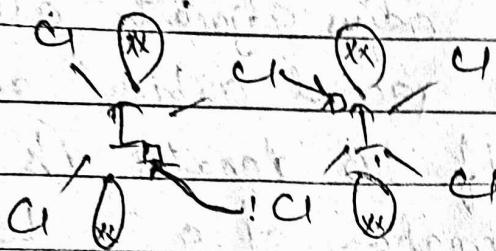
The bridge bonds the
are above and below
the plane and
the plane of
the molecule

$\rightarrow BeCl_2$ Solid phase (polymer)



Ex^m, Al_2Cl_6 , Fe_2Cl_6 , Ti_2Cl_6

Ti_2Cl_6

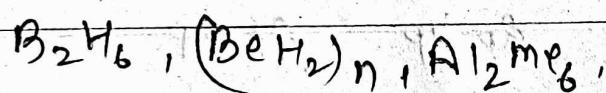


Here lone pairs are above and below the plane of molecule and bridge atoms are in the plane of molecule.

TRICK

$3\text{C} - 2\text{e}^-$ bond

→ Side atom = H, me, etc.



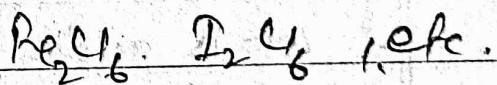
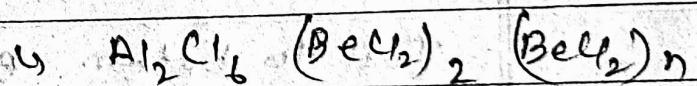
→ e^- deficient,

→ banana bond

$3\text{C} - 4\text{e}^-$

side atom Cl (which have lone pair)

e^- rich.



* Co-ordinate bond :- When bond is formed by donation of e^- pair. $\text{A}^* \rightarrow \text{B}$

(electron rich - non - metal)

donor atom

acceptor atom

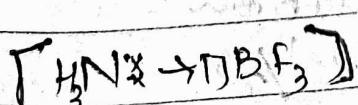
Ex^m $\text{H}_3\text{N}^* \rightarrow \text{D}_2\text{H}$

lone's base

Lewis Acid.

$\Rightarrow \text{NH}_4^+$ type of

Co-ordinate bond is a ~~part~~ covalent bond.

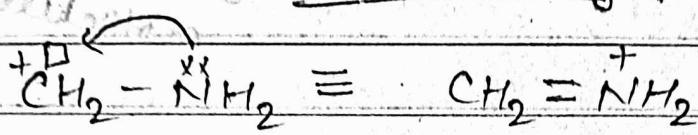


1:1 complex
OR
Adduct

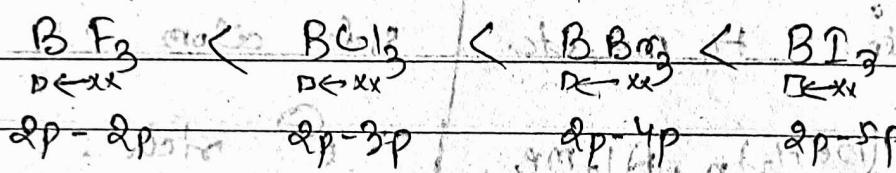
* Back Bonding



A and B are two adjacent atoms which are bonded. A have vacant orbital and B have lone pair. Then if B donates its lone pair to A, resulting formation of a π bond, is known as Back bonding.



* Compare Lewis acidic strength.



Top to bottom B.B. strength \downarrow est. (size diff. \uparrow)

Lewis acidic strength T to bottom \uparrow est.

* Vanderwaal forces (weaks intermolecular forces)

① Ion-Dipole (ion-polar molecule)

② Ion-Induced dipole (ion-non-polar molecule)

③ Dipole-Dipole (persione forces) (two polar mol.) Strong

④ Dipole-Induced dipole (Debye forces) (polar-nonpolar)

⑤ Induced dipole-Induced dipole. (London/Dispersion force)

* Relation b/w interaction energy & distance (two non-polar) force relation.

D.P-D.P (stationary)

D.P-D.P (rotating)

D.P-D.D.P

I.D.P-I.D.P

$$E \propto \frac{1}{r^3}$$

$r = \text{dist. b/w}$
 two molecules

$$E \propto \frac{1}{r^6}$$

$$E \propto \frac{1}{r^6}$$

$$E \propto \frac{1}{r^6}$$

* B.P. and M.P. in covalent compound.

B.P. & M.P. \propto H-bond
 \propto van der waal force
 \propto Molecular weight.

① Hydrides of 14th group

B.P. & M.P. $\text{CH}_4 < \text{SiH}_4 < \text{GeH}_4 < \text{SnH}_4 < \text{PbH}_4$

② Hydrides of 15th group

B.P. $\text{PH}_3 < \text{AsH}_3 < \text{NH}_3 < \text{SbH}_3 < \text{BiH}_3$

↓
due to intermolecular H-bond.

M.P. $\text{PH}_3 < \text{AsH}_3 < \text{SbH}_3 < \text{BiH}_3 < \text{NH}_3$ (due to effective packing)

③ Hydrides of 16th group

M.P. & B.P. $\text{H}_2\text{S} < \text{H}_2\text{Se} < \text{H}_2\text{Te} < \text{H}_2\text{Po} < \text{H}_2\text{O}$

④ Hydrides of 17th group

B.P. & M.P. $\text{HCl} < \text{HBr} < \text{HI} < \text{HF}$ (due to H-bond)

⑤ B.P. $\text{C}_2\text{H}_6 > \text{CH}_4$ surface area $\uparrow \rightarrow$ H-H force \uparrow
 $\text{B.P.} \uparrow$

M.P. : Symmetrical \rightarrow Unsymmetrical alkane.

Metallic Bonding : The force exist b/w +ively charged kernels and mobile e^- . \star [Nucleus + inner core e^-]
 \hookrightarrow kernels.

① Metallic bond strength \propto no. of mobile e^-
alkali $<$ Alkaline earth metals.

② Temp $\uparrow \rightarrow$ metallic conductivity \downarrow