

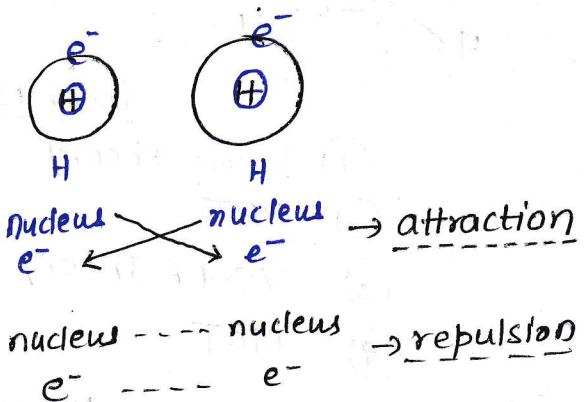
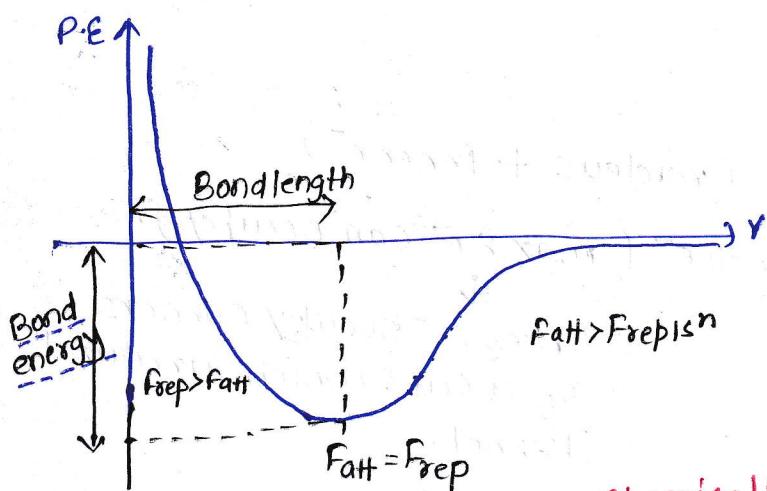
Chemical Bond

- atom → do not exist → independently (except Noble gas)
- Group of atom → exist → molecule → The attractive force which holds atom, ions together in a molecule is called chemical bond.
- attraction ↑ Energy ↓ Stability ↑ Strength of bond.

cause of chemical bond

- ① Tendency to acquire min energy →

- ② octet rule



Chemical bond

- ~~Octet Rule~~

Strong bond (interatomic bond)

- Covalent
 - ionic
 - metallic
 - coordinate
- (200-400 KJ/mol)

weak bond (intermolecular bond)

- Vanderwall (2-8 KJ/mol)
- H-Bond (8-40 KJ/mol)

(2-40 KJ/mol)

Covalent bond

- Sharing of electrons to complete its octet. (H → duplet)
- Bond formed b/w nonmetal & non-metal.



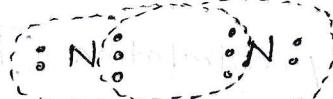
Single cov bond

$2e^-$ share



double cov bond

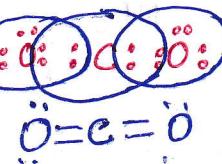
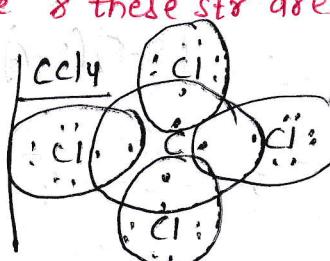
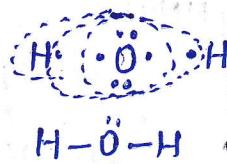
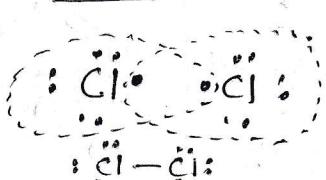
$4e^-$ share



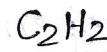
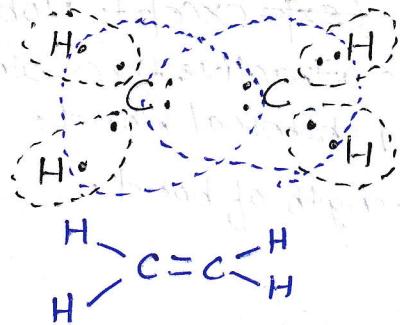
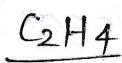
Triple cov bond

$6e^-$ share

The dots represent Valence e^- & these str are k/a Lewis str.



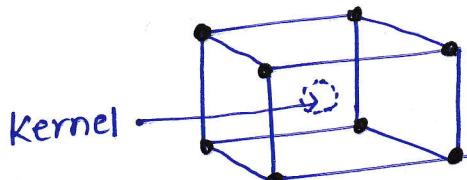
Chemical Bond



Theories of Covalent bond

- ① Lewis theory
- ② VBT
- ③ VSEPR Theory
- ④ MOT

Lewis theory: atom \rightarrow Kernel (nucleus + inner e^-)
 \rightarrow outer shell e^- (max 8 e^- can occupy).



These 8 e^- occupy corners of a cube which surround kernel.

Octet Rule → given by Lewis & Kossel

also k/a electronic theory of chemical bond

Atoms combine to complete its octet in their outermost orbit

Sharing of e^-
e.g. $\text{Cl}_2, \text{N}_2, \text{O}_2$ etc

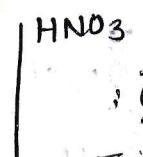
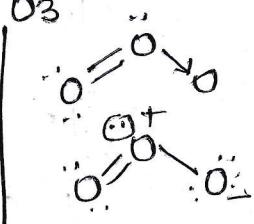
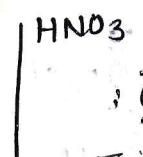
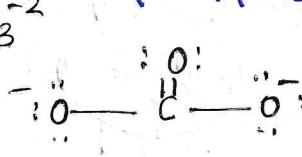
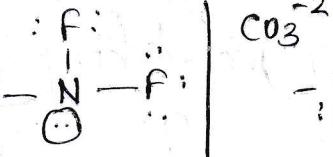
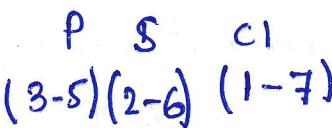
complete transfer of e^-
e.g. $\text{NaCl}, \text{CaCl}_2, \text{MgO}$ etc.

Rules to draw Lewis dot str

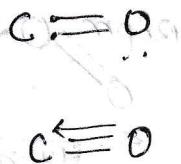
• Identify central atom (least EN atom, less in number, high atomic number, large size)
 forms multiple bond.

• arrange Li Be B C N O F ~~N~~

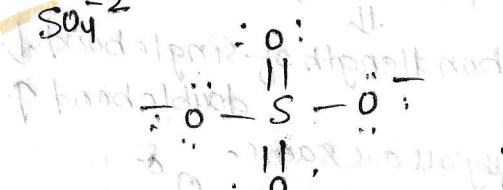
| | | | | | | |
|---|---|---|---|---|---|---|
| 1 | 2 | 3 | 4 | 3 | 2 | 1 |
|---|---|---|---|---|---|---|



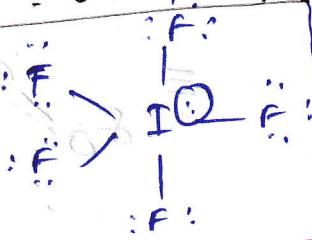
CO



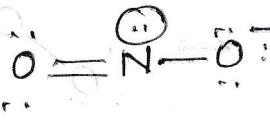
SO_4^{2-}



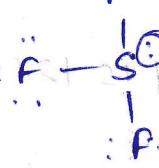
IF_5



NO_2^-



SF_4

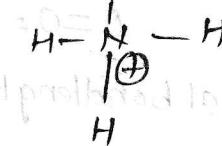
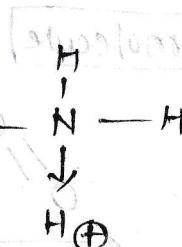


NO



odd e⁻ molecule
octet Rule fails

NH_4^+



NO_2

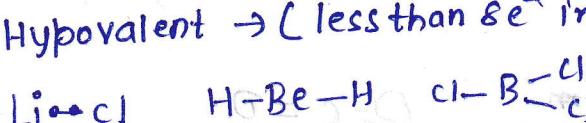


$\frac{7e^-}{11}$

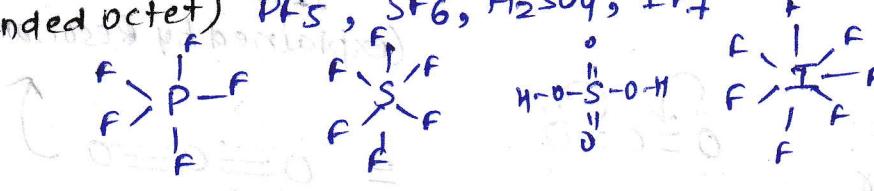
odd e⁻
octet rule fails

Drawbacks of Lewis theory

- Could not explain str of any compd.
- Could _____ B.L, B.O, B.E etc.
- Hypervalent \rightarrow (less than 8 e⁻ in valence shell)



- Hypervalent \rightarrow (more than 8 e⁻ in valence shell)
(expanded octet) PF_5 , SF_6 , H_2SO_4 , IF_7



Odd electron species

NO & NO_2 & ClO_2

- Compounds of noble gases are also formed
 XeF_2 , KrF_2 , XeOF_2 , XeO_3 , XeOF_4 , XeF_6

Formal charge

- Formal charge of an individual atom is formal charge

$$\text{F.C.} = \text{valence e}^- - (\text{bonds} + \text{dots})$$

$$\textcircled{1} \quad \text{O}_3 \quad \rightarrow 6 - (3+2) = \textcircled{1}$$

$$\rightarrow 6 - (1+6) = \textcircled{1}$$

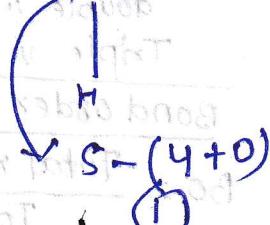
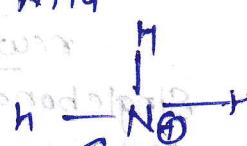
$$6 - (2+4) = \textcircled{0}$$

CN^-



$$4 - (3+5) = \textcircled{0}$$

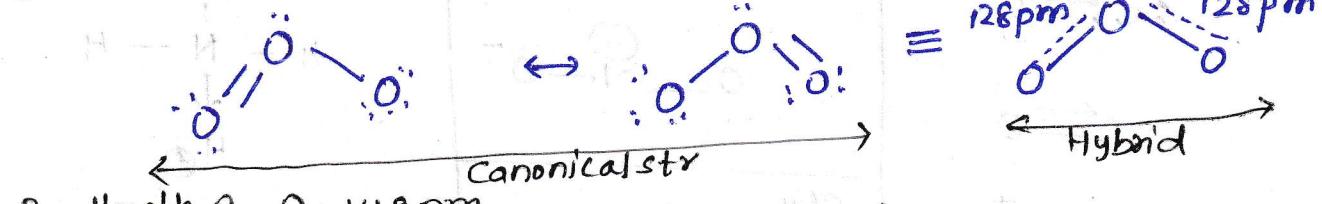
NH_4^+



①

O₃ molecule

Resonance



Bond length O-O = 148 pm

“ ” O=O = 121 pm

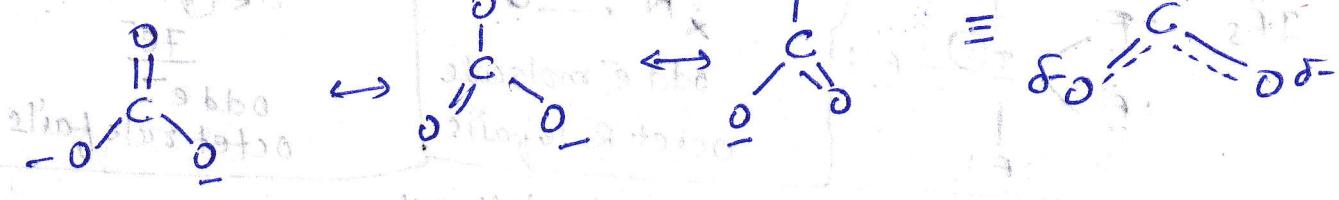
actual bond length of O₃ o bond is 128 pm (explained by Reso)

\downarrow
bond length of single bond
 \uparrow
double bond

②

CO₃²⁻ ion

Two types of carbon oxygen bond
C=O and C-O but experimentally all are same.



③

CO₂ molecule

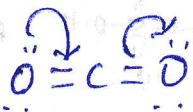
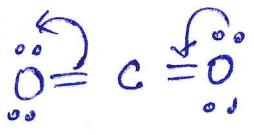


actual C=O \rightarrow 121 pm

“ ” C \equiv O \rightarrow 110 pm

but in CO₂ $\text{O}=\text{C}\equiv\text{O}$ \downarrow \downarrow 115 pm

(explained by Resonance)



- Resonance hybrid is real str. & Canonical forms are hypothetical.
- Resonance hybrid has less energy than energy of any single canonical str so, Resonance stabilizes the molecule.

Bond order

neutral molecule

Single bond \Rightarrow BO = 1

double " $= 2$

Triple " $= 3$

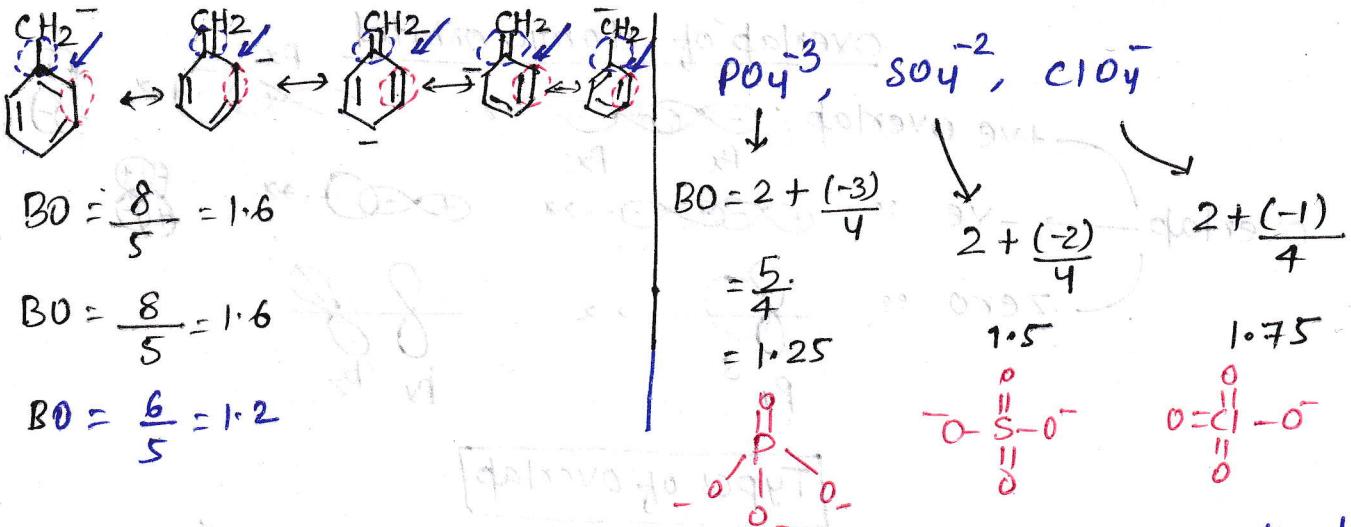
Bond order in Resonance str

BO = $\frac{\text{Total no of bonds b/w two atoms in all str}}{\text{Total no of Rs}}$

Radical (charged)

PO₄³⁻, SO₄²⁻, ClO₄⁻

BO = Valency of + charge of com
SA no of SA



Bond order & Bond strength & Bond En & Stab $\propto \frac{1}{\text{Bond length}}$

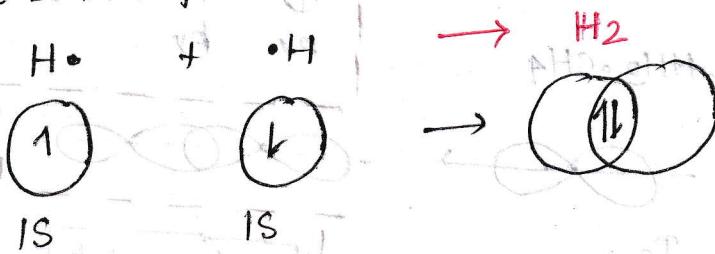
If Bond order = +ve. Integer \rightarrow compd is stable
= 0/-ive \rightarrow Unstable

= fraction (greater than 1) \rightarrow compd should isomerism

$= +\frac{1}{2}$ \rightarrow Unstable but exist $\begin{matrix} H_2^+ \\ He^+ \end{matrix}$

Valence Bond Theory \rightarrow Overlapping Theory
 \rightarrow Hybridization Theory

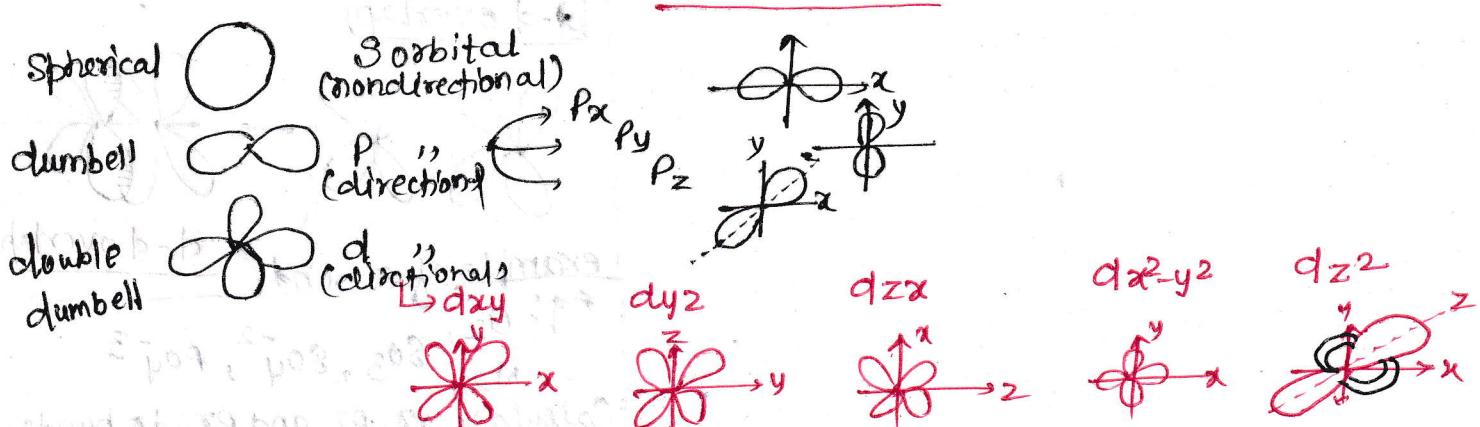
Covalent bond is formed when atomic orbital combine.



- overlapping orbital \rightarrow half filled
 \rightarrow must have opposite spin
- Extent of overlap should be less than 50% (in case of more than 50% repulsion b/w both nucleus occurs)

Extent of overlap or Strength of bond or stability of bond energy $\propto \frac{1}{\text{Bond length}}$.

Atomic orbital

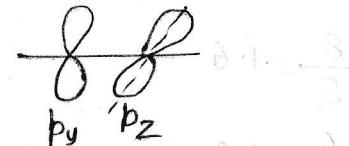


overlap of atomic orbital

+ve overlap -  $p_x \quad p_x$

Overlap \rightarrow -ve "  $p_x \quad p_x$

zero "  $p_y \quad p_y$

 $p_z \quad p_z$

Types of overlap

axial or head-on

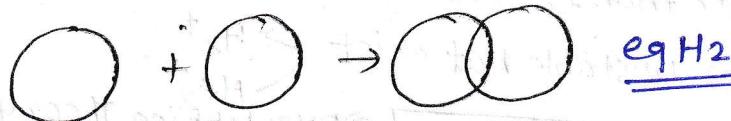
Sigma bond

S-S overlap

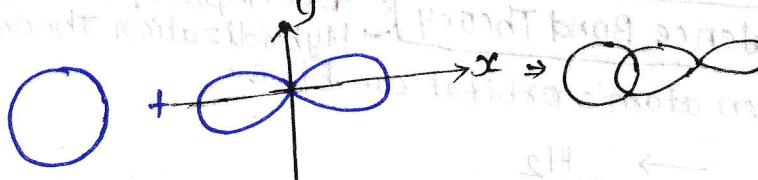
S-P "

P-P "

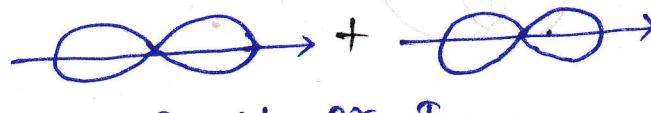
S-S



S-P



P-P



$$p_x + p_x \rightarrow x = INA$$

$$p_y + p_y \rightarrow y = INA$$

$$p_z + p_z \rightarrow z = INA$$

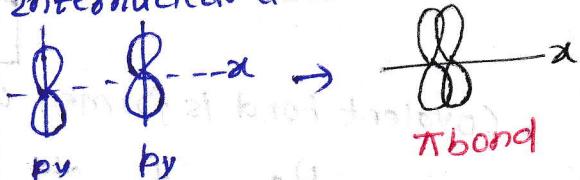
} axial/head on
overlap

parallel or
sideways
Pi bond

P-P overlap, P-d overlap
d-d overlap

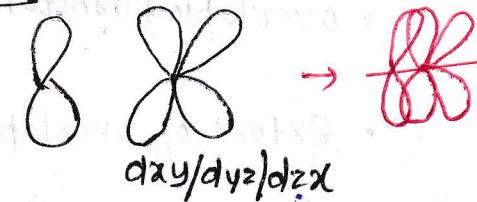
P-P overlap \rightarrow overlap not on internuclear axis

overlap is above or below internuclear axis.

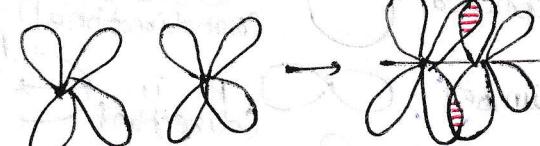


$p_z + p_z \rightarrow$ axis xory
 $p_z + p_y \rightarrow$ axis yorz

P-d overlap



d-d overlap

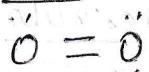


example of π bond:

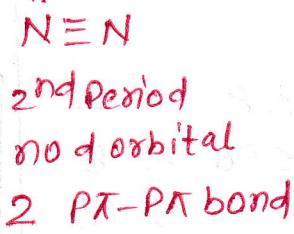
(eg: $O_2, SO_3, SO_4^{2-}, PO_4^{3-}$)

calculate $P\pi-P\pi$ and $P\pi-d\pi$ bonds

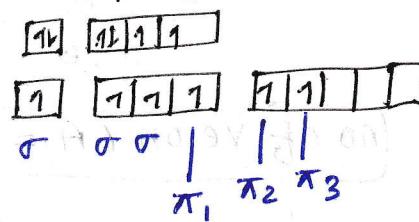
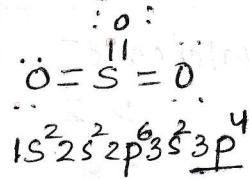
O₂



N₂

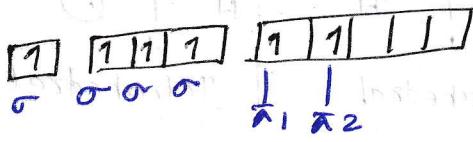
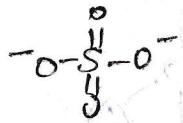


SO₃



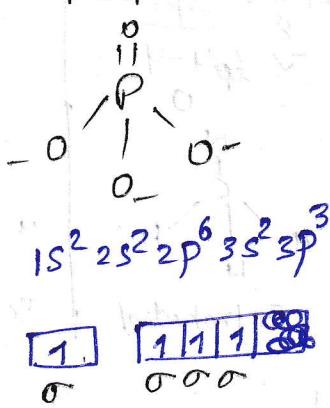
(4)

SO₄⁻²



2 d π -P π

PO₄⁻³



1 d π -P π

Hybridization

orbital of slightly different energy

→ orbital of equivalent energy & shape

• for bond formation

① intermixing of orbital → Hybridization → Hybrid orbital.

② overlapping of orbital

(i) • num of hybrid orbital = no of atomic orbital intermixed.

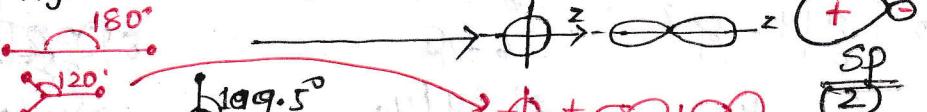
(ii) • Hybrid orbital are more effective in forming stable bond than atomic orbital.
↳ bcoz hybrid orbitals have less Energy than Pure orbital.

(iii) • It is a theoretical concept

(iv) • electron do not participate in hybridization only orbitals participate.

(v) • Hybrid orbital arrange themselves in such a geometry in which repulsion is min. (Hybridization gives geometry).

H₂O 2 SP ^{Geo} Linear



3 Sp² Trigonal planar



4 Sp³ Tetrahedral



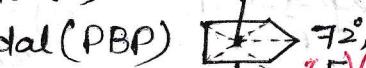
5 Sp^{3d} Trigonal bipyramidal (TBP)



6 Sp^{3d}₂ Octahedral



7 Sp^{3d}₃ Pentagonal bipyramidal (PBP)



(not in NCERT)

↳ 3 ⁺ 3 Sp²

Calculation of Hybridization

① Sum of Valence e⁻

| | | |
|-------|------|------------------|
| ② Sve | > 2 | $\leq 8 \div 2$ |
| | > 8 | $\leq 56 \div 8$ |
| | > 56 | $\div 18$ |

Steric number rule

Hybridsⁿ = no of σ bond
+ no of lone pair

③ no of ve on LA = HO - SA

Bcl₂

$$2 + 2 \times 7$$

$$8 | 16 | 2 - SP \\ \frac{16}{x} \quad 1p = 2 - 2 = 0$$

Cl - Be - Cl
Linear

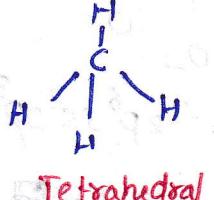
Bcl₃

$$3 + 3 \times 7 \\ 8 | 24 | 3 - SP^2 \\ \frac{24}{x} \quad 1p = 3 - 3 = 0$$

Cl
Cl
B - Cl
T.P.

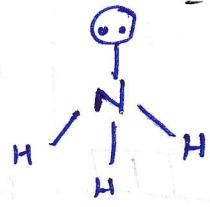
CH₄

$$4 + 1 \times 4 \\ 2 | 8 | 4 - SP^3 \\ \frac{8}{x} \quad 1p = 4 - 4 = 0$$



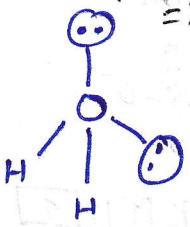
NH₃

$$5 + 1 \times 3 \\ 2 | 8 | 4 - SP^3 \\ \frac{8}{x} \quad 1p = 4 - 3 = 1$$



H₂O

$$6 + 1 \times 2 \\ 2 | 8 | 4 - SP \\ \frac{8}{x} \quad 1p = 4 - 2 = 2$$



Pcl₅

$$5 + 7 \times 5$$

$$8 | 40 | 5 - SP^3d \\ \frac{40}{x} \quad 1p = 5 - 5 = 0$$

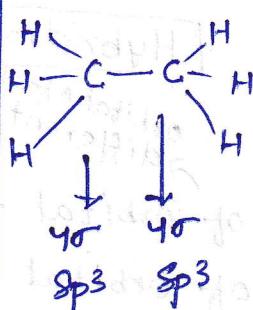
Cl
Cl
Cl
Cl
P - Cl
Cl
Cl
F

SF₆

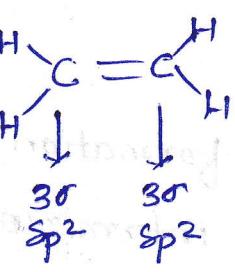
$$6 + 7 \times 6 \\ 8 | 48 | 6 - SP^3d^2 \\ \frac{48}{x} \quad 1p = 6 - 6 = 0$$

F
F
S
F
F
F
F

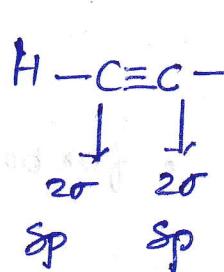
C₂H₆



C₂H₄



C₂H₂



TBP

Octahedral

NH₄⁺

$$5 + 4 - 1$$

$$2 | 8 | 4 - SP^3$$

H
N
H
H

SF₆

$$6 + 7 \times 4$$

$$6 + 28$$

$$8 | 34 | 4 > 5$$

$$2 | 2 | 1 - SP^3d$$

$$\frac{4}{x} \quad 1p = 5 - 4 = 1$$

SO₄⁻²

$$6 + 6 \times 4 + 2$$

~~8 | 32 | 4~~

~~2 | 4 | 2~~

~~2 | 2 | 0~~

~~8 | 32 | 4~~

~~2 | 4 | 0~~

$$5 + 6 \times 3 + 1$$

$$5 + 18 + 1$$

$$8 | 24 | 3 - SP^2$$

$$\frac{24}{x} \quad 1p = 0$$

$$O^- N O^-$$

Top planar

~~8 | 32 | 4~~

~~2 | 4 | 0~~

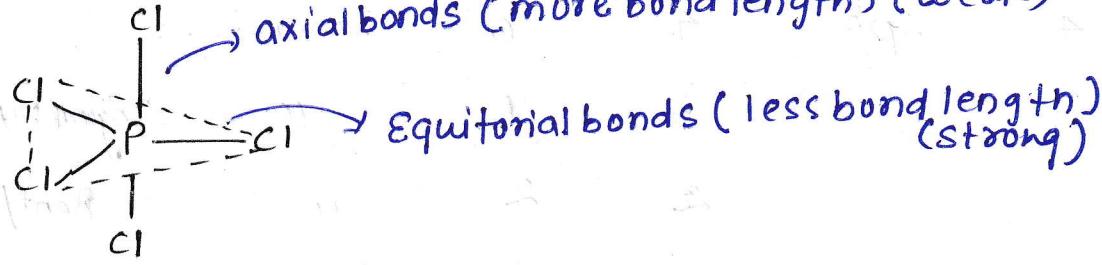
~~8 | 32 | 4~~

~~2 | 4 | 0~~

~~8 | 32 | 4~~

~~2 | 4 | 0~~

Pcl₅

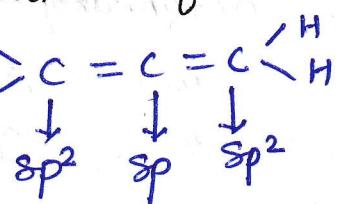


PF₅

all equatorial & axial bonds have same bond length.
(bcz F is highly En)

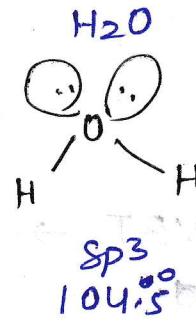
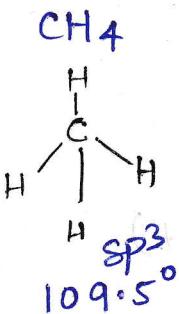
sp³d \rightarrow dz² orbital hybridizes.

Ques Call hybridization of all carbon in allene (C₂H₄)



VSEPR Theory

Valence Shell electron pair repulsion Theory



lp-lp repulsn > lp-bp repulsn > bp-bp repulsn

Geometry

\rightarrow depends on lp & bp both.

Shape

\rightarrow depends on bp only.

Geo - Tetrahedral

Shape - "

Hybrid orbital

• 2

sp

• 3

sp²

BP
2

LP
0

Geo
Linear

Tetrahedral

pyramidal

Tetrahedral

Bent/angular

Shape

Linear BeCl₂, HgCl₂,

T.P BF₃,

bent/angular

SO₂/O₃

3
2

0
1

T.P

T.P



4 sp^3 4 O Tetra

(4 p-orbital broad band) 3 broaded 1 twips \rightarrow " " \rightarrow

Tetrahedral CH₄, NH₄⁺

2 2 .. Bent/angular H₂O

5 sp^3d

5 O T-B-P

T-B-P PCl₅

4 sp³d¹ 3 broaded 1 twips \rightarrow " " \rightarrow Seesaw SF₄

3 2 .. T-shape ClF₃

2 3 .. Linear xEF₂

6 sp^3d^2

6 O Octahedral Octahedral

SF₆

5 1 .. Square pyramidal BrF₅

4 2 .. Square planar xEf₄

7 sp^3d^3

7 O P.B.P P.B.P

IF₇

Capped octahedral xEf₆
distorted octahedral

MOT

Molecular Orbital Theory

Alc VBT \rightarrow O₂ \rightarrow diamagnetic X

O₂ \rightarrow paramagnetic (experiment)

↑
MOT

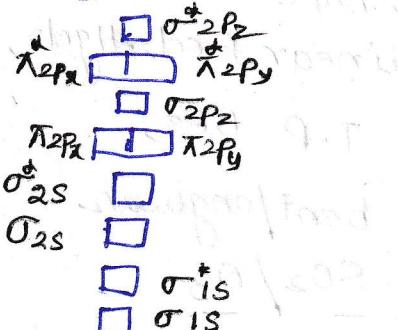
• Two atomic orbital first overlap ~~overlapping~~ to form 2 MO.

AO + AO \rightarrow BMO (Bonding molecular orbital) \rightarrow less Energy
(constructive interference)

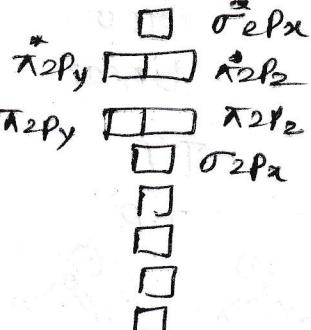
ABMO (Antibonding molecular orbital) \rightarrow more En
(destructive interference)

• In MOT, e⁻ are filled alc to Aufbau, Pauli, Hund's Rule.

$\leq 14e^- - 2121$



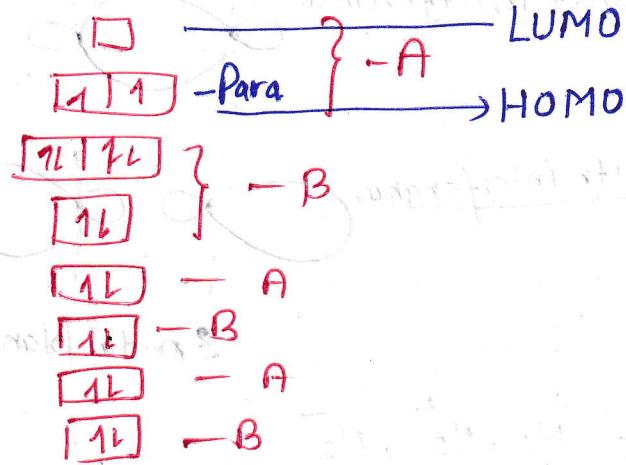
$> 14e^- - 1221$ lower energy orbital is filled first



first all orbitals of same energy will be singly occupied then pairing starts

9n orbital electrons are opposite.

O₂ - 16 e⁻ 122)



$$BO = 2$$

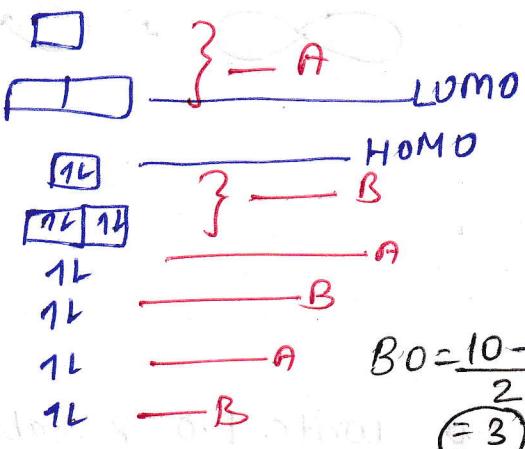
$$BO = \frac{10-6}{2}$$

$$= \frac{4}{2}$$

$$= 2$$

$$\text{Bondorder} = \frac{N_b - N_a}{2}$$

N₂ - 14 e⁻ 212)



BO ~~for species due to~~ (Unstable)
~~BO~~ = 1 - Sing bond
 2 - double "
 3 - Triple "

$$BO = 10 - 4 = 3$$

- BO ↑ double bond character ↑ Stab ↑ B.L ↓
- unpaired e⁻ → para
- no unpaired e⁻ → dia

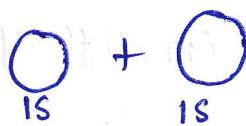
| | |
|----|-----|
| 10 | 1 |
| 11 | 1.5 |
| 12 | 2 |
| 13 | 2.5 |
| 14 | 3 |
| 15 | 2.5 |
| 16 | 2 |
| 17 | 1.5 |
| 18 | 1 |

all odd e⁻ → para } except 10 & 16
 all even → dia }

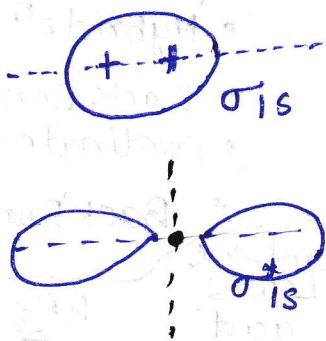
if Bo of 2 molecule is same then
 more antibonding e⁻ → less stable
 less " " " " → more stable

Nodes & Nodal plane

↓ Point at which probability of finding electron is zero.
 ↓ Plane at which finding electron is zero.



→ Constructive interference → + + σ_{1s}
 → destructive " " → - - σ_{1s}

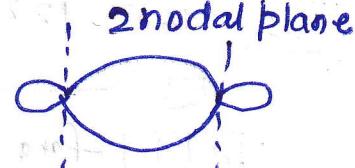


1 node
1 nodal plane

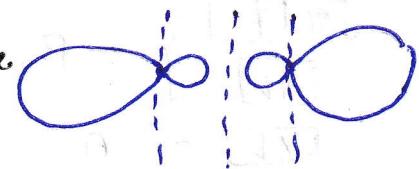
P



Const Interference



destr interference



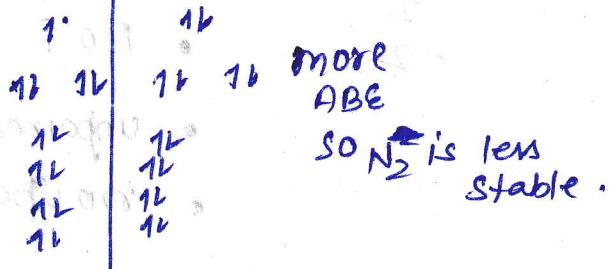
3 nodal plane.

Ques write B.O & stability order of N_2, N_2^+, N_2^-

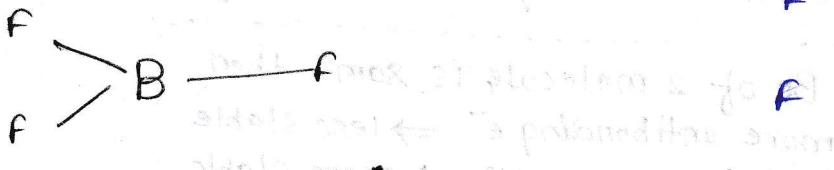
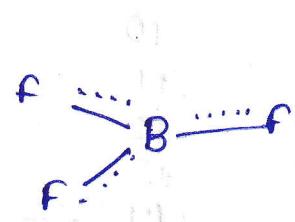
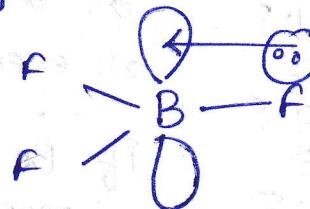
\downarrow \downarrow \downarrow
 3 2.5 2.5

(more ABE \rightarrow less stable)

| Stability |
|-----------------------|
| $N_2 > N_2^+ > N_2^-$ |



Back Bond



$$\rightarrow 1.30 \text{ \AA} \text{ (experiment)}$$

- It takes place b/w 2 bonded atom
- one atom should have lone pair to donate & other atom should have vacant orbital.
- Bond strength always \uparrow & Bond length always \downarrow
- Hybridization, Bond angle may or may not change.
- Back Bond (lone pair donated to vacant orbital in same species).
- Coordinate (" " " " " " " in diff species).

• Back Bond $\propto \frac{1}{\text{Lewis acid character}}$

Component
Lewis acid
character

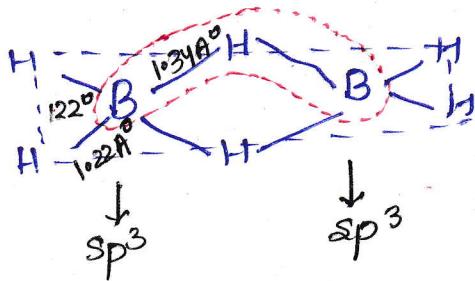
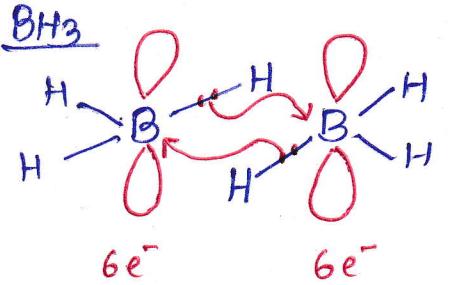


Banana Bond Bonding in e⁻ deficient comp

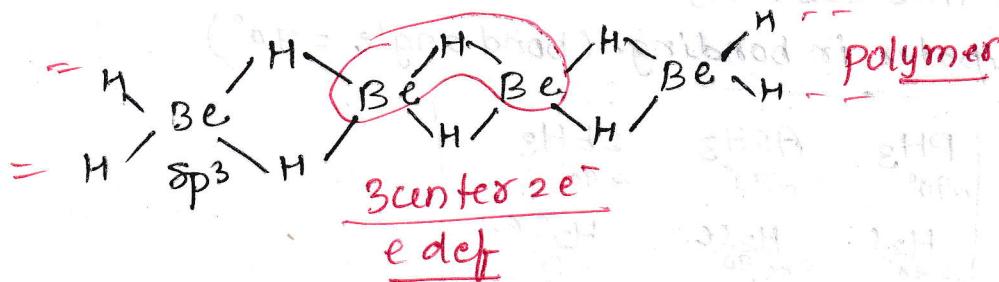
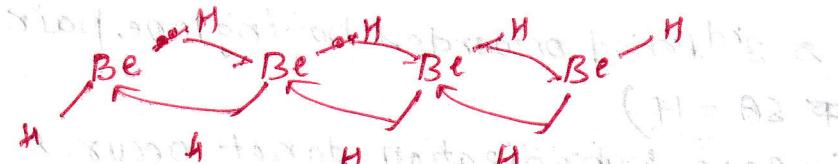
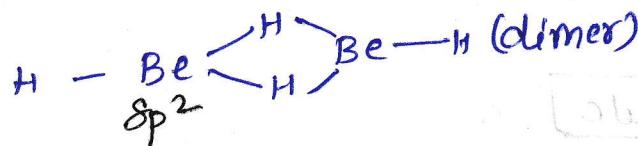
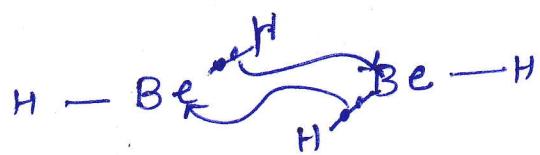
BH₃, BeH₂, BCl₃, AlCl₃

3 center 2 e⁻ 3 center 4 e⁻
(e⁻ def)

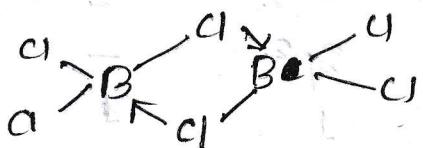
- Boron momentarily completes its octet.
- Hybridization \rightarrow $sp^2 \rightarrow sp^3$
- $B_2H_6 \rightarrow$ ~~planar~~ \rightarrow (e⁻ def)
- $4H \rightarrow$ same plane] \rightarrow 6 atom in one plane
 $2B \rightarrow$ "]
- 2 bridging hydrogen \rightarrow out of plane
- $B_2H_6 \rightarrow$ non planar



BeH₂ \rightarrow sp



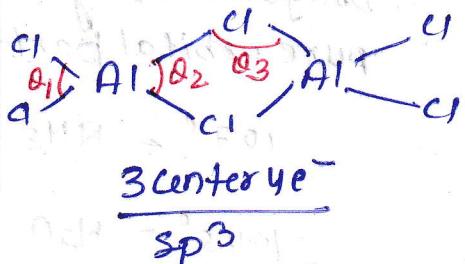
BCl₃



3 center 4 e⁻

- Stable
- non planar
- not e⁻ def

AlCl₃



$O_1 > O_2 > O_3$

Comparison of Bond parameter

① BA \propto % S

$SP - 50\% S$

$SP^2 - 33\% S$

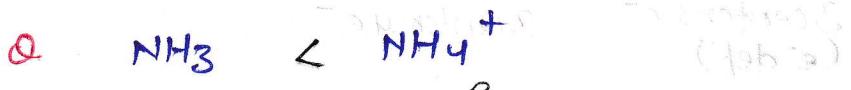
$SP^3 - 25\% S$

② BA \propto $\frac{1}{\text{ion radius}}$

③ BA \propto En of CA

④ BA \propto $\frac{1}{\text{en of SA}}$

Compare Bond angle :-



DRAGO Rule

- CA \rightarrow G15, G16, G17 & 3rd Period onwards & having lone pair

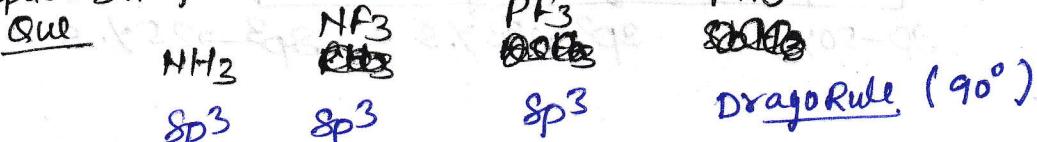
- SA \rightarrow En ≤ 2.5 (~~esp~~ SA = H)

Drago rule says in this case hybridization don't occur & pure orbital participates in bonding. (bond angle = 90°)



| PH ₃ $\sim 90^\circ$ | ASH ₃ $\sim 90^\circ$ | SbH ₃ $\sim 90^\circ$ |
|-------------------------------------|--------------------------------------|--------------------------------------|
| H ₂ S $\sim 90^\circ$ | H ₂ Se $\sim 90^\circ$ | H ₂ Te $\sim 90^\circ$ |
| : Br-H $\sim 90^\circ$ | : F-H $\sim 90^\circ$ | |

Compare B.Angle



Lonep
catom

—
—

PH₃
Drago Rule (90°)

$\text{NH}_3, \text{NF}_3 > \text{PF}_3 > \text{PH}_3$

B.Ang $\frac{1}{\text{En of SA}}$ $\leftarrow \text{NH}_3 > \text{NF}_3 > \text{PF}_3 > \text{PH}_3$

(8)

Bond length

$$\text{① } \text{B.L} \propto \frac{1}{4\text{EN}}$$

$$\text{② } \text{B.L} \propto \frac{1}{\text{Bo}}$$

③ Resonance & Bond ^{Back} \downarrow Bond length

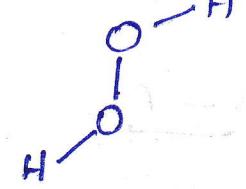
one) compare Bond length $\text{O}_2 < \text{O}_3 < \text{H}_2\text{O}_2$



Complete
DB



Partial
DB



Single
Bond

Bond energy

$$\text{B.E} \propto \Delta\text{EN}$$

$$\text{B.E} \propto \text{B.O}$$

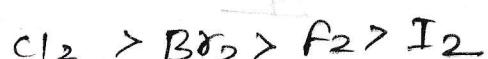
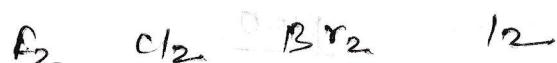
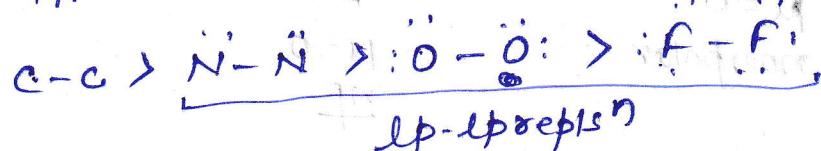
$$\text{B.E} \propto \frac{1}{\text{size of atom}}$$

one) compare Bond energy

one

one

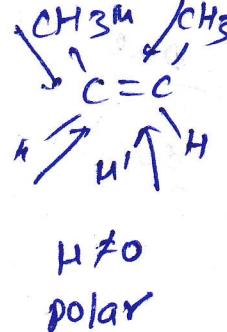
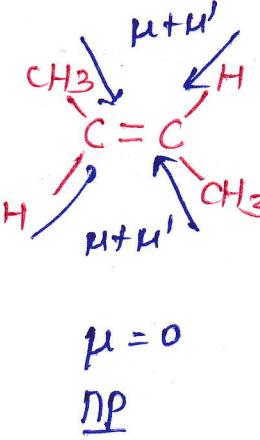
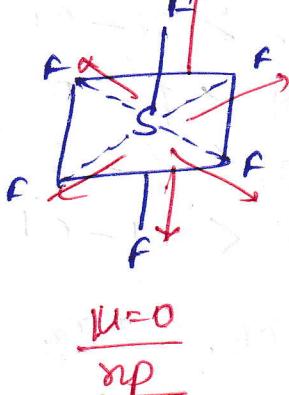
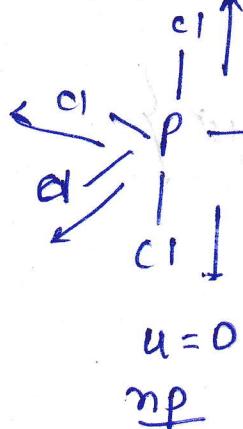
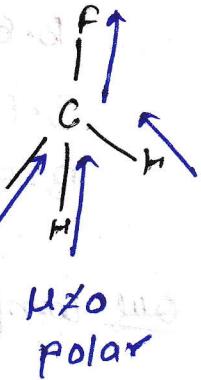
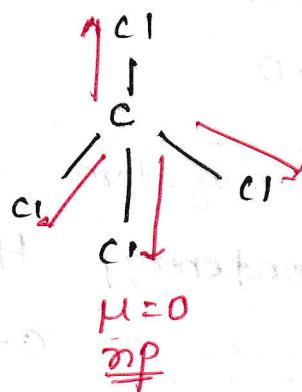
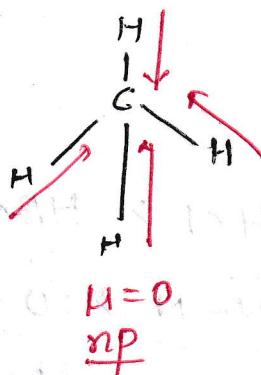
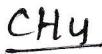
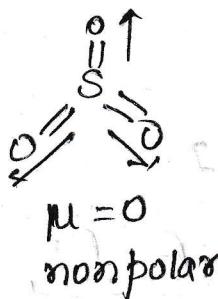
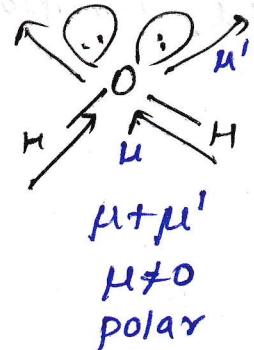
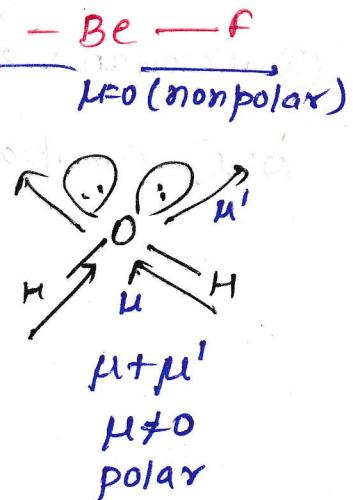
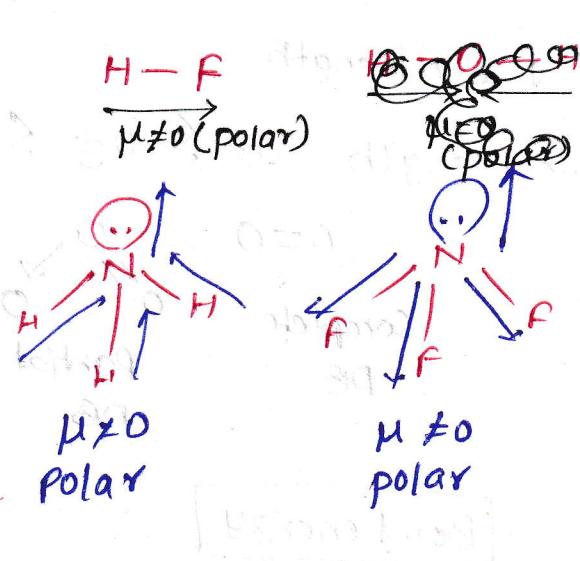
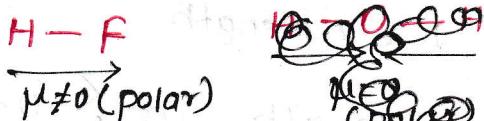
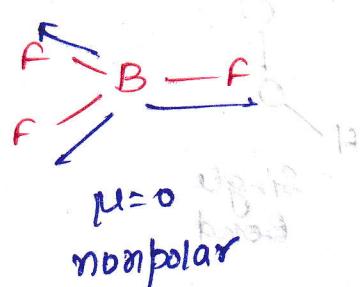
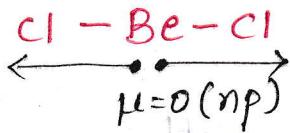
one



Dipole moment → Tells about polarity of a molecule.

- Draw arrow (less En to more En) (Atom to lone pair)

- equal & opposite vector cancel each other.



Secondary bond \rightarrow (2-40 kJ/mol)

- ① vanderwall force of attraction (2-8 kJ/mol)
- ② Hydrogen Bond (8-40 kJ/mole)

Vanderwall

vanderwall \propto At wt/mol wt

Vanderwall \propto B.P.

Ques $F_2, Cl_2 \rightarrow$ gas, $Br_2 \rightarrow$ liq, $I_2 \rightarrow$ solid. Explain

$$P_2 < Cl_2 < Br_2 < I_2 \text{ (mwt) (VFOA)}$$

① Ion-dipole $Na^+ \cdots H_2O$

② dipole-dipole $+ - \cdots + -$

③ Ion-induced dipole $Na^+ \cdots Br_2 \rightarrow Na^+ \cdots Br_2^+$

④ dipole induced dipole $+ - \cdots + - \rightarrow + -$

⑤ Instantaneous induced dipole
(London dispersion force)

Ques HCl has more B.P than H_2 . Why??

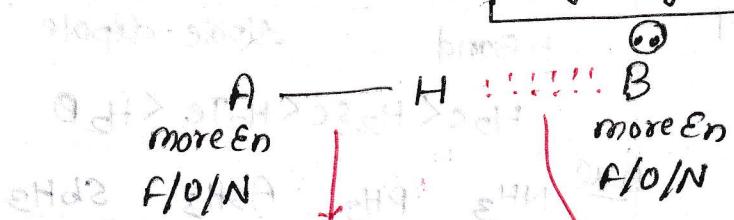
\downarrow
dipole-dipole

\downarrow
L.D.F

Ques compare BP $H_2, Cl_2, N_2, Br_2, C_2, P_2$

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

Hydrogen Bond



\rightarrow H Bond
& 40 kJ/mol

Strength of H Bond \propto En of A

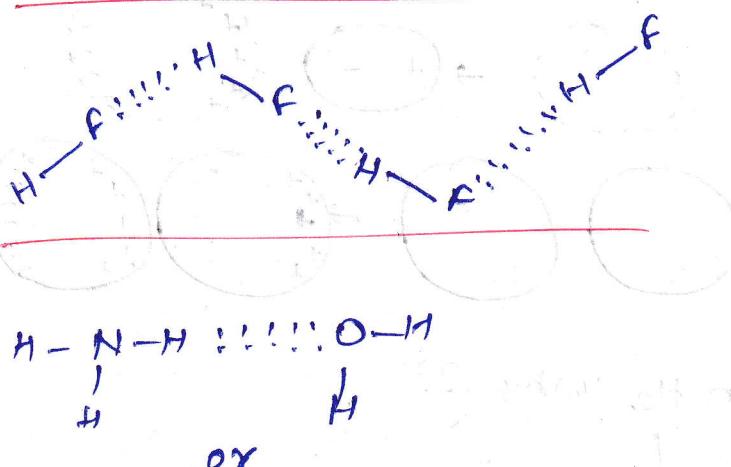
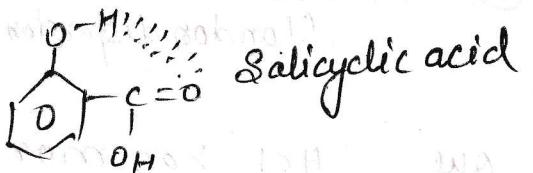
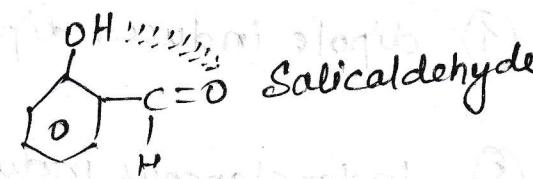
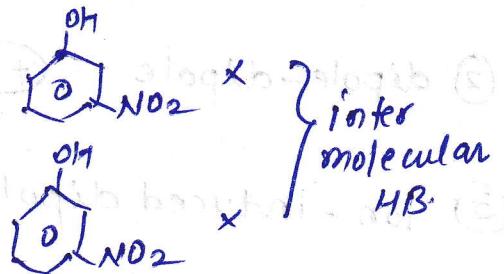
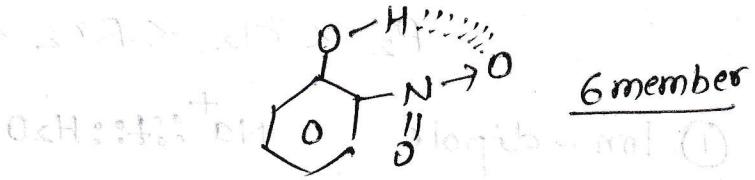
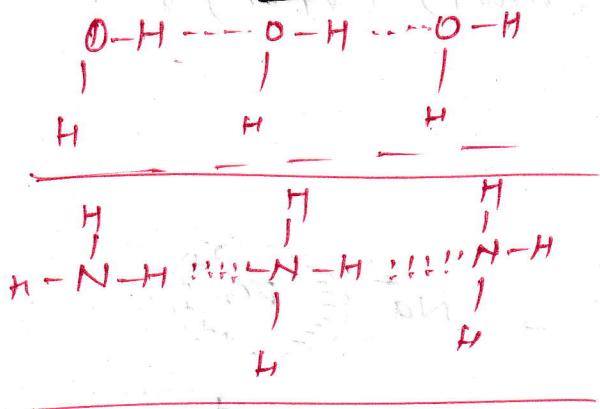
$$\propto \frac{1}{\text{En of B}}$$

Ques Compare H-Bond Strength

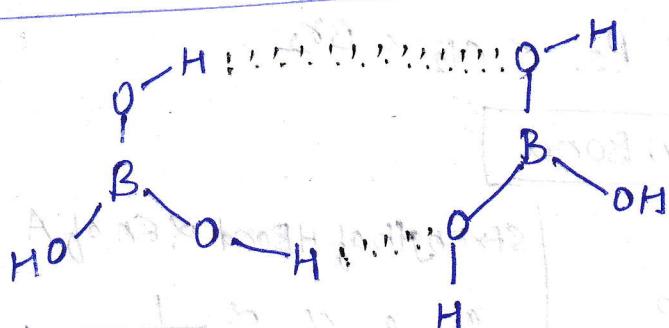
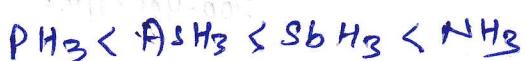
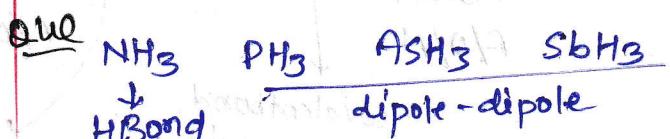
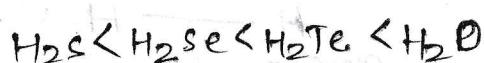
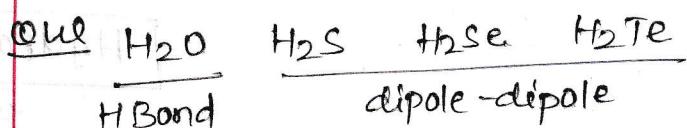
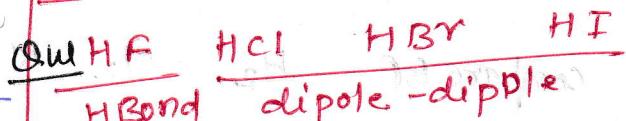


HBond

Intermolecular \rightarrow Intermolecular \rightarrow Intermolecular \rightarrow Intermolecular



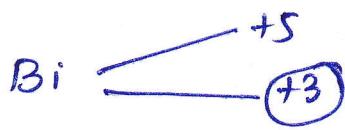
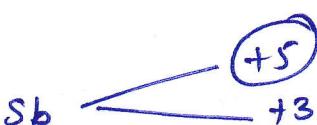
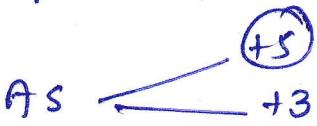
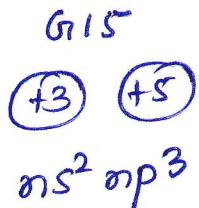
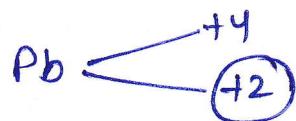
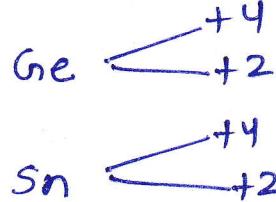
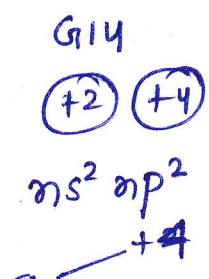
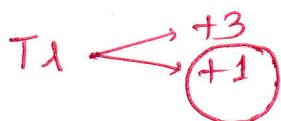
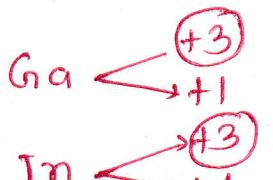
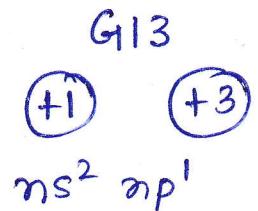
Compare B.P.



$\text{H}_3\text{BO}_3 \rightarrow$ exists as dimer

Inert pair effect

- only applicable in G₁-13, G₁-14, G₁-15.
- as we move down the group, z_{eff} increases (poor shielding of d & f electrons)
- ~~$z_{eff} \uparrow$ so, ns^2 electrons are not excited for reaction (attraction of nucleus) & become inert.~~ $\rightarrow 8 ns^2 e^-$ become inert



- as we move down the group stab of lower O.S. ↑.