PRACHAND NEET

ONE SHOT

Inorganic Chemistry

[redacted]

Chemical Bonding and Molecular Structure

By – Hitesh Sopra Sir

Topics *to be covered*

1. Chemical Bonding One Shot

PRACHAND SERIES

TELEGRAM CHANNEL

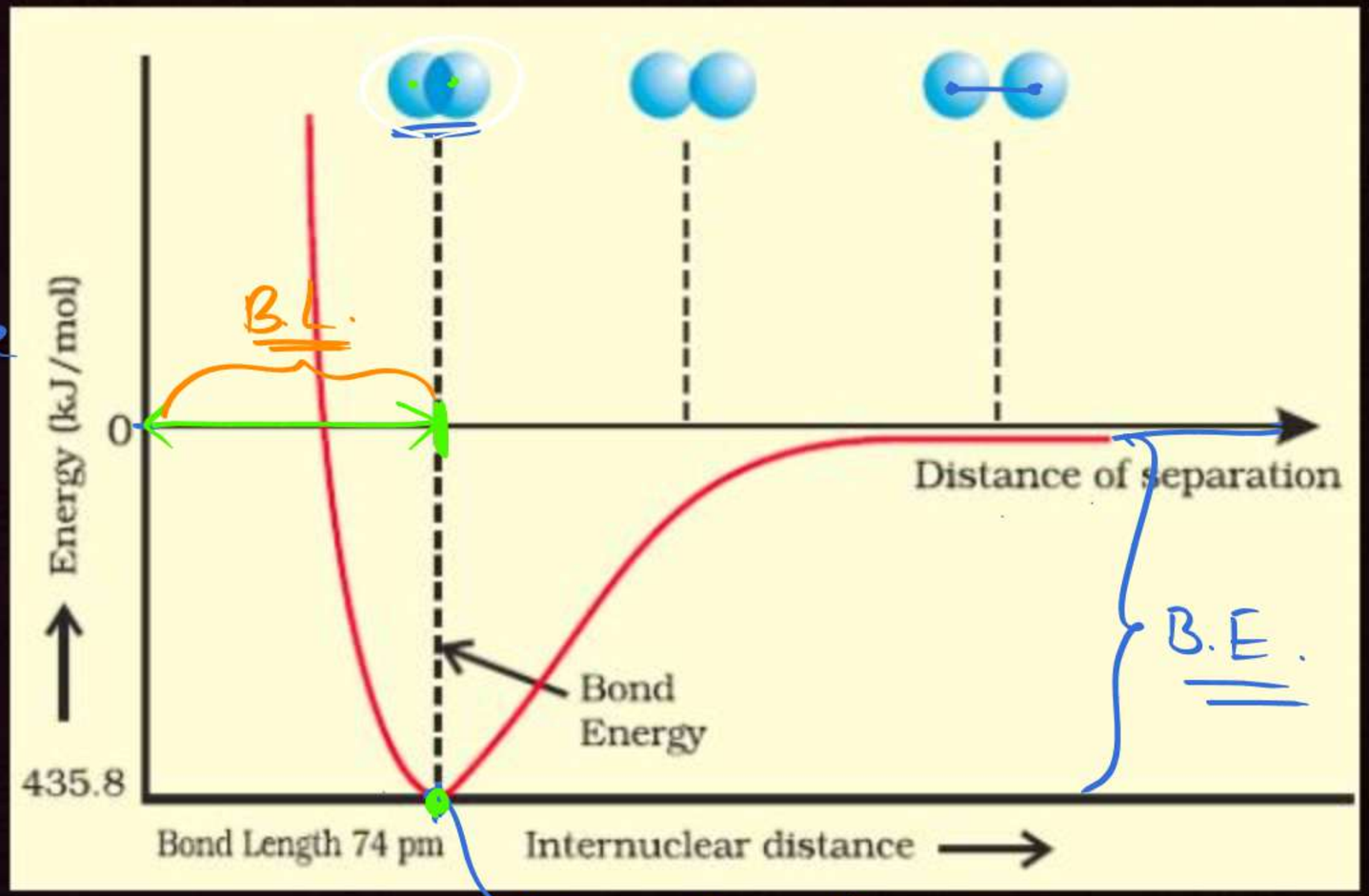
Introduction

Bond → force of attⁿ.

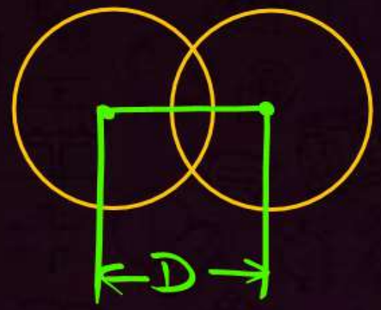
B.E. ↑ Bond Enthalpy = -ve (Exothermic)

* Note :→ Bond formⁿ. is always an exothermic proces.

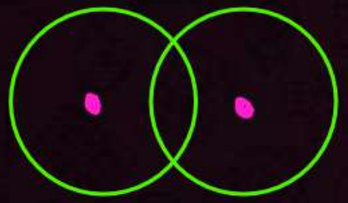
* Energy Concept *



* **Bond length** :→ Internuclear distance of 2 bonded atoms.



Stability $\propto \frac{1}{\text{Energy}}$



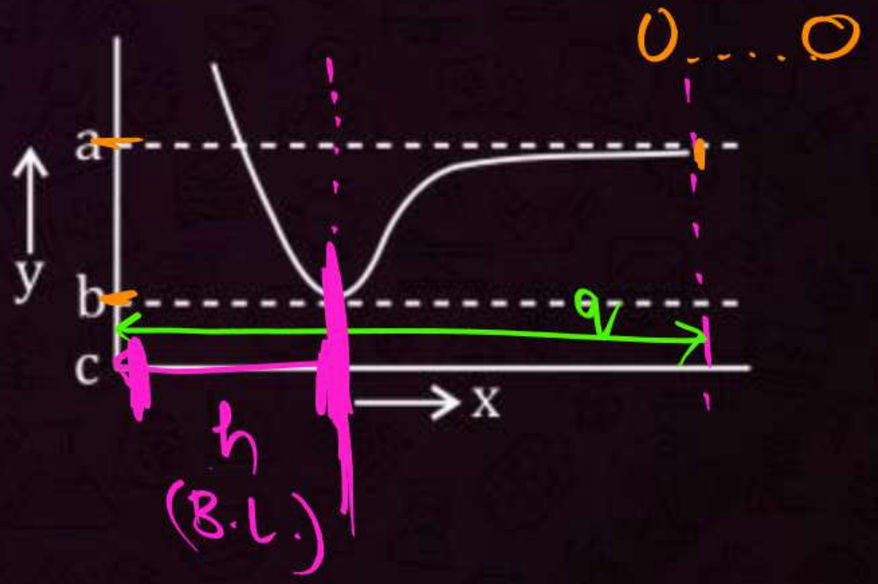
QUESTION [2020-Covid]

The potential energy (y) curve for H_2 formation as a function of internuclear distance (x) of the H atoms is shown below.

The bond energy of H_2 is;

H — H

B.L. = ? (h)



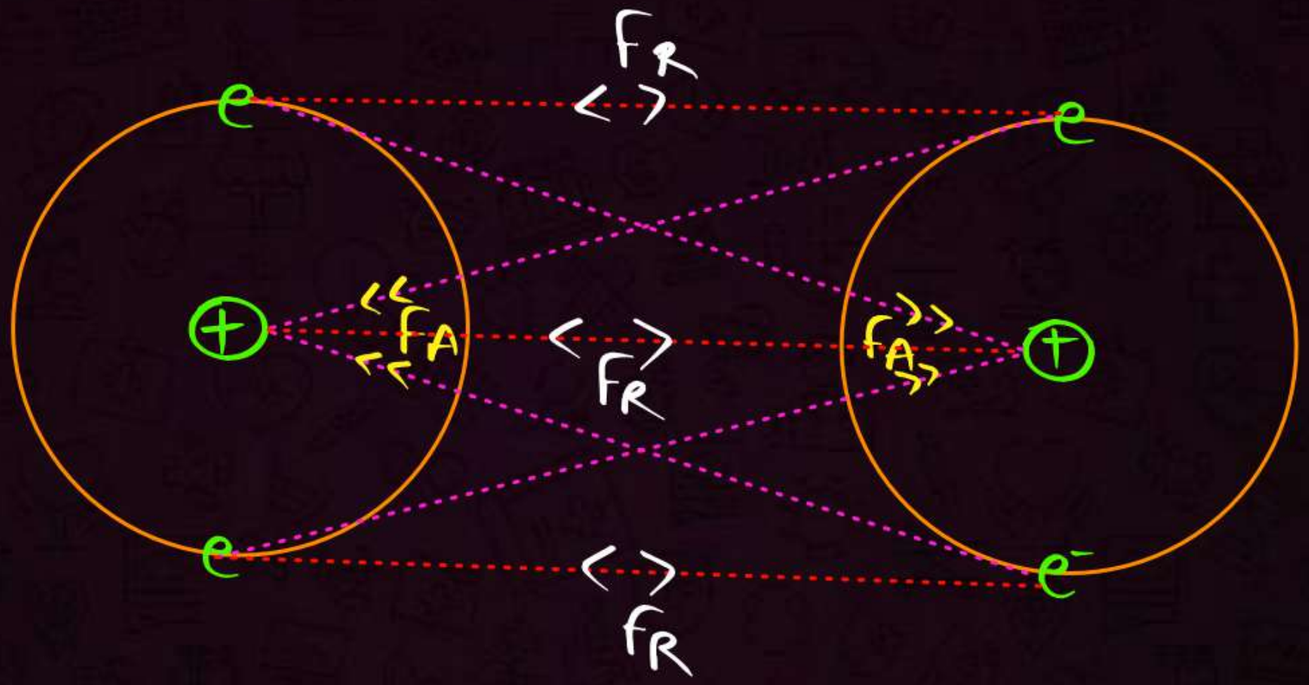
(A) $\frac{(c-a)}{2}$

(B) $\frac{(b-a)}{2}$

(C) $(c-a)$

(D) $(b-a)$ ✓

* Condition of Bond formation :



Condition of Bond form^n :→ $F_A > F_R$

Octet Theory

* Kossel & Lewis Concept *

He → $1s^2$ (2)

Ne → $1s^2 2s^2 2p^6$ (2, 8) → octet

Ar → (2, 8, 8)

Kr → (2, 8, 18, 8)

Xe → (2, 8, 18, 18, 8)

Rn →

Og →

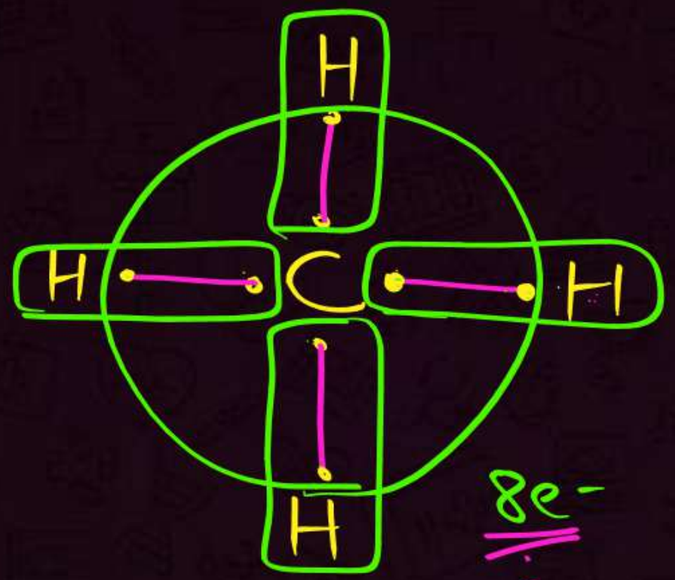
Valence shell
(outermost shell)

* Octet Rule :→ Atoms combine to complete their resp. octet / Duplet.

CH_4

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

$4e^-$



$8e^-$

$NaCl$

$Na \rightarrow 1s^2 2s^2 2p^6 3s^1$

$Na^+ \rightarrow 1s^2 2s^2 2p^6$

$8e^-$

$Cl \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^5$

$Cl^- \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6$

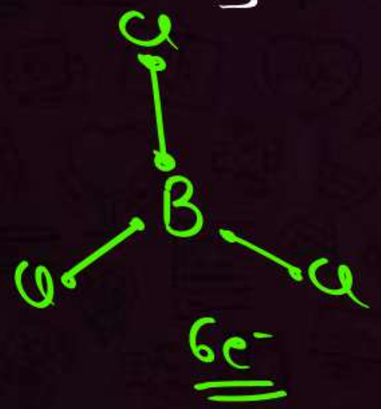
$8e^-$

Na^+ Cl^-

Exceptions of Octet Theory

I) Hypovalent / e^- difficient species :↓

Ex: BCl_3



$BeCl_2$

*Note :→ All hypovalent species are Lewis Acids.

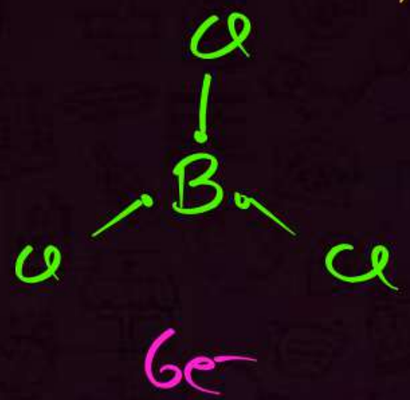
*NEET Q W.O.F can act as Lewis Acid?

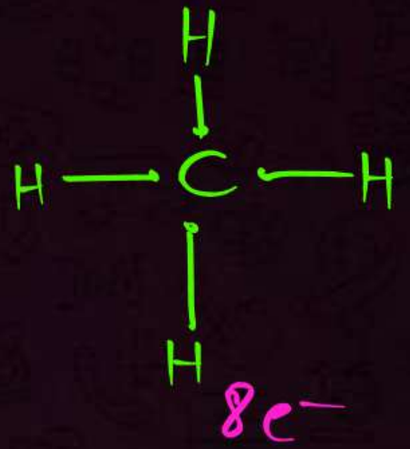
① BCl_3

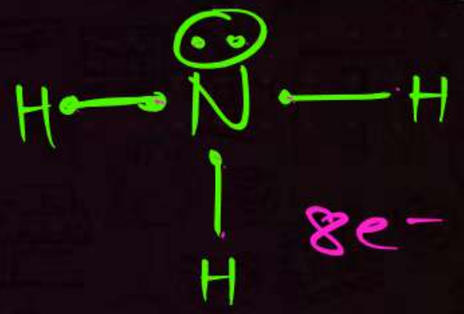
② CH_4

③ NH_3

④ All





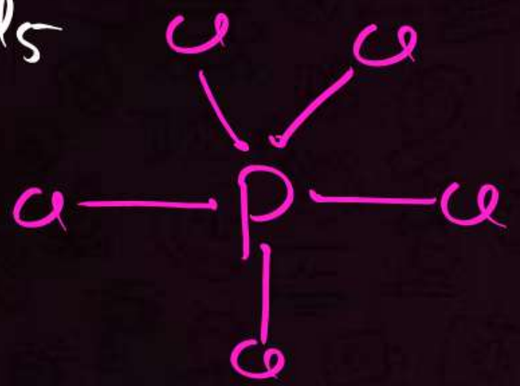


P-block

B_2H_6 ⇒ e^- defficient (Hypovalent)

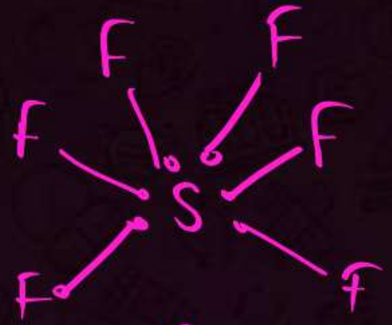
II) e⁻ rich / hypervalent species :→ more than 8e⁻

PCl_5



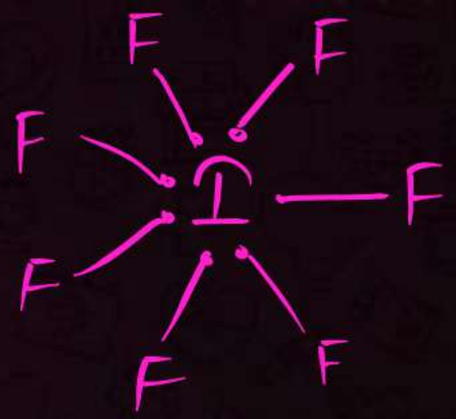
10e⁻

SF_6



12 e⁻

IF_7



14e⁻

* Note :→ e⁻ rich species can also behave as Lewis Acid

~~e⁻ rich species are Lewis Base~~

$PF_5 + F^{\ominus} \rightarrow [PF_6]^-$

(III) Odd e^- species :→ odd no. of e^-.

Ex : NO ⇒ 7 + 8 = 15 e^-

NO_2 ⇒ 7 + 8 + 8 = 23 e^-

ClO_2 ⇒ 17 + 8 + 8 = 33 e^-

ClO_3 ⇒ 17 + 8 + 8 + 8 = 41 e^-

7 e^-

*Note :→ All odd e^- species are paramagnetic → feels attⁿ. in magnetic field.

Q W.O.F. specie attract Magnetic field lines towards itself?

Paramagnetic

(1) N_2O_5

(2) N_2O_4

(3) NO_2 → odd e⁻ species ⇒ Paramagnetic

(4) N_2O_3

Q W.O.F. specie Repel Magnetic field lines?

Diamagnetic (All e⁻ paired)

① NO

② CO

③ NO_2

④ CO_2

(IV) Pseudo Noble gas configuration :-

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

$n=1$ $n=2$ $n=3$

$18e^-$ → Pseudo Noble gas Config

Ex: Zn^{+2}, Cu^{+}, Ag^{+}, Ga^{+3}, Ge^{+3} etc...

Cd^{+2}

Hg^{+2}

V) Non Inert gas Configuration :→

(18+2) e⁻

Inert Pair effect (P-block)

Ex: Tl^{+1}, Pb^{+2}, Bi^{+3} etc.---

$Tl^{+1} \rightarrow [Xe]\ 4f^{14}\ 5d^{10}\ 6s^2\ 6p^1$

$5s^2\ 5p^6$

(18 + 9)

n = 6

VI) 9 to 17 e⁻ config →

Formal Charge

The formal charge of an atom in a polyatomic molecule or ion may be defined as the difference between the number of valence electrons of that atom in an isolated or free state and the number of electrons assigned to that atom in the Lewis structure. It is expressed as :

$$F.C. = V - B - U$$

V → no. of valence e⁻ in free state

B → No. of Bonds

U → No. of Unshared e⁻

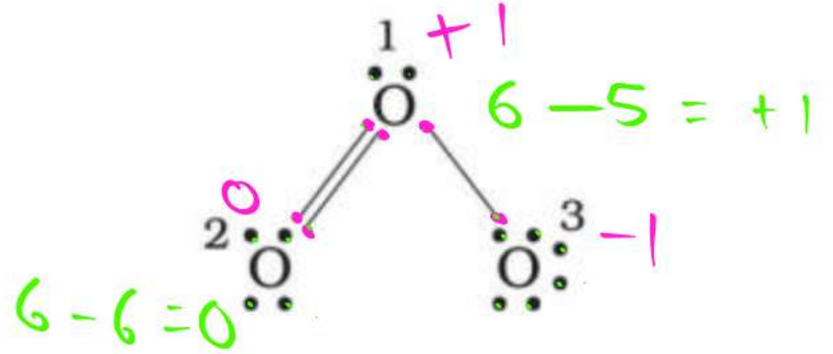
Formal charge (F.C.) on an atom in a Lewis structure (F.C.)	=

[total number of valence electrons in the free atom] − [total number of non bonding (lone pair) electrons] − (1/2) [total number of bonding (shared) electrons]

NCERT CORNER

$O \rightarrow 1s^2\ 2s^2\ 2p^4$ ($6e^-$)

Let us consider the ozone molecule (O_3). The Lewis structure of O_3 may be drawn as:



The atoms have been numbered as 1, 2 and 3. The formal charge on:

$F.C. = V - B - U$

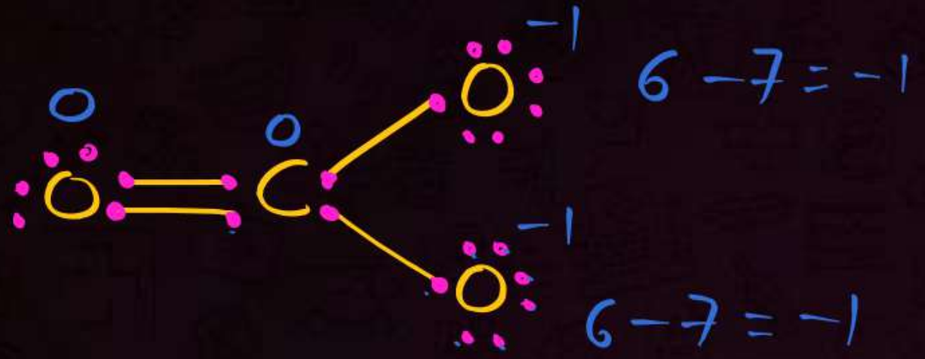
(O_1) $\quad 6 - 3 - 2 = +1$

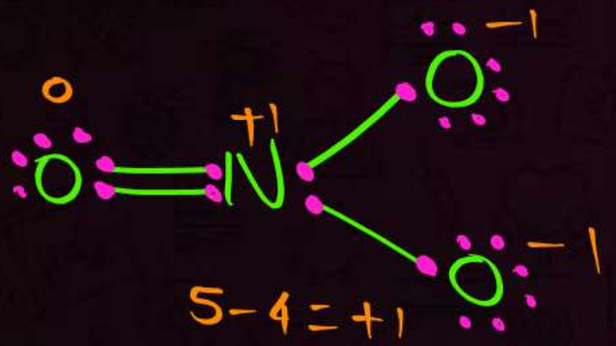
$F.C. = 6 - 2 - 4 = 0$ $\quad (O_2)$

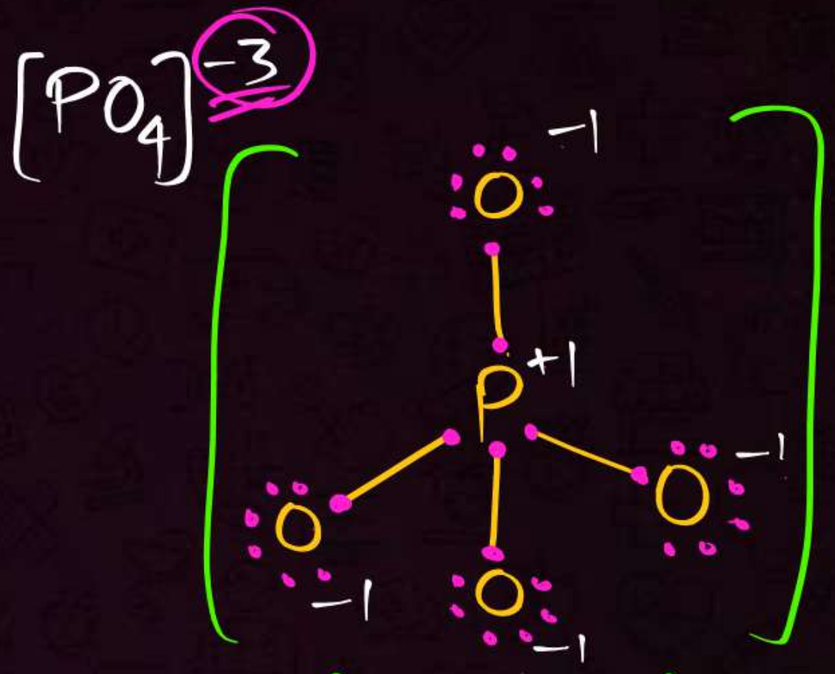
$F.C. = 6 - 1 - 6 = -1$ $\quad (O_3)$

Hitsick ⇒ जितने e^- होते हैं − जितने e^- हैं

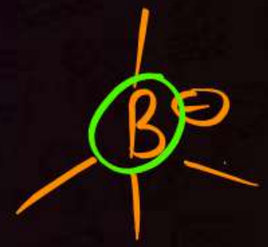
$6 - 7 = -1$

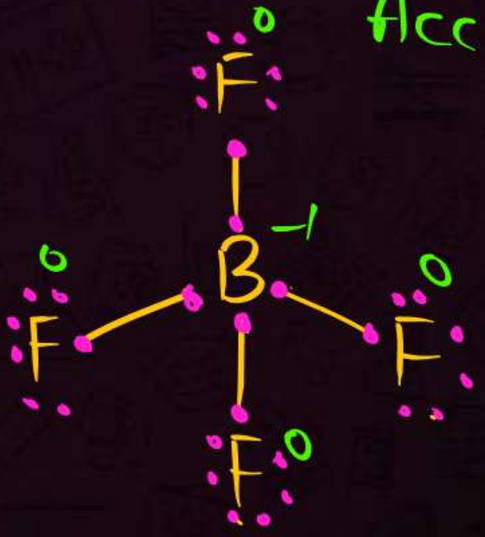


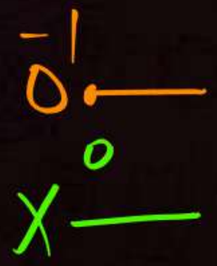




Acc. to octet theory

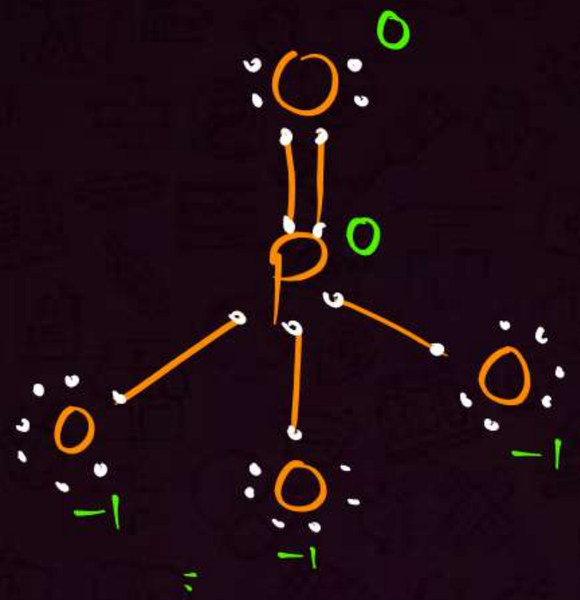






7 - 7 = 0

3 - 4 = -1



$[PO_4^{-3}]$

Q F.C. present on central oxygen atom in O_3 ?

(1) +1

(2) -1

(3) $-\frac{1}{2}$

(4) $+\frac{1}{2}$

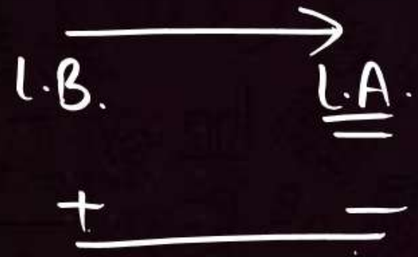
Coordinate Bond

Covalent Bond $A—B$

Coordinate Bond $A \leftarrow B$

$$L.A. \leftarrow L.B$$

⇒ Dative Bond

⇒ 

L.A.

① Vacant orbital must be present

② e⁻ defficient / hypovalent specie

③ No stearic repulsion

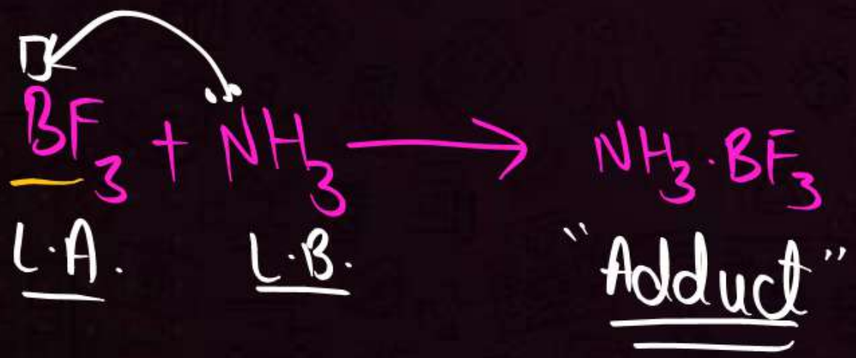
L.B

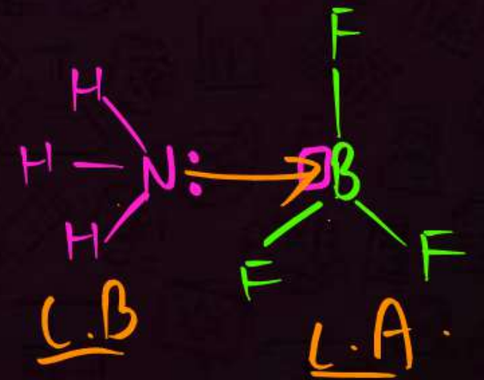
① e⁻ pair must be present.

② octet complete.

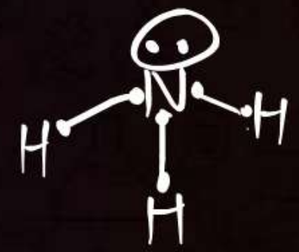
③ Tendency to donate.

L.A.
BF_3





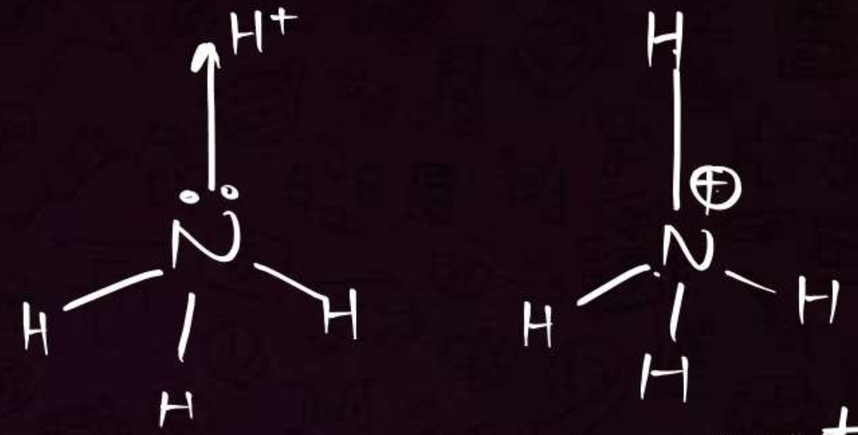
$\ddot{N}H_3$
L.B.



Protonation Rxn :→

$NH_3 + H^+ \longrightarrow NH_4^+$

L.B. L.A.



NH_4^+

$H_7O_3^+$

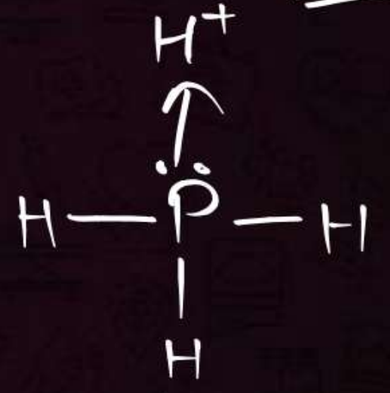
H_2O

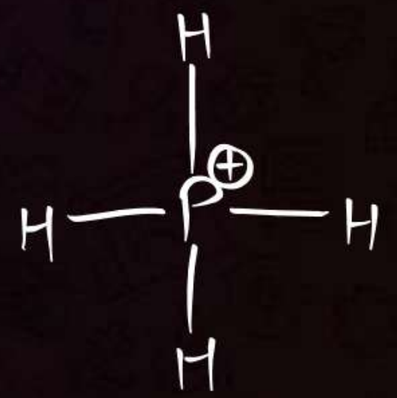
H_3O^+

H_2O

$H_5O_2^+$

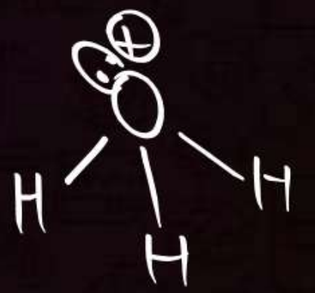
$PH_3 + H^+ \longrightarrow PH_4^+$



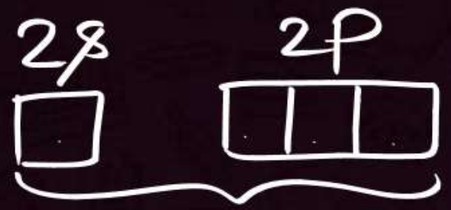


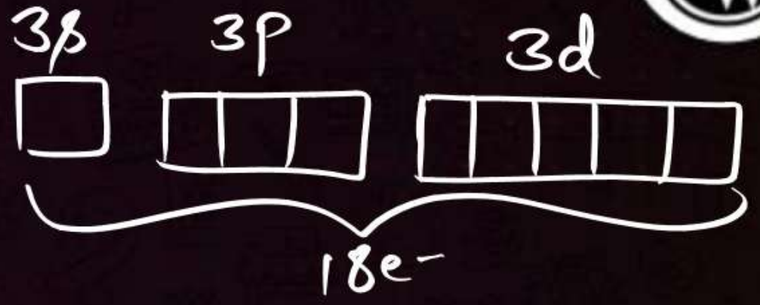
PH_4^+

$H_2O + H^+ \longrightarrow H_3O^+$



*Note :→ 2nd period Elements can't exceed their octet





~~BF_6^{-3}~~ ~~NF_5~~

~~OF_6~~ ~~NCl_5~~

~~CF_6^{-2}~~

~~$ClF_5^{\ominus}$~~

Q W.O.f. does not exist?

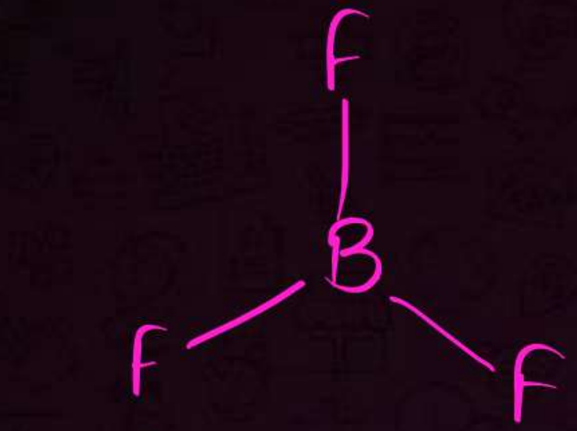
(A) NCl_3

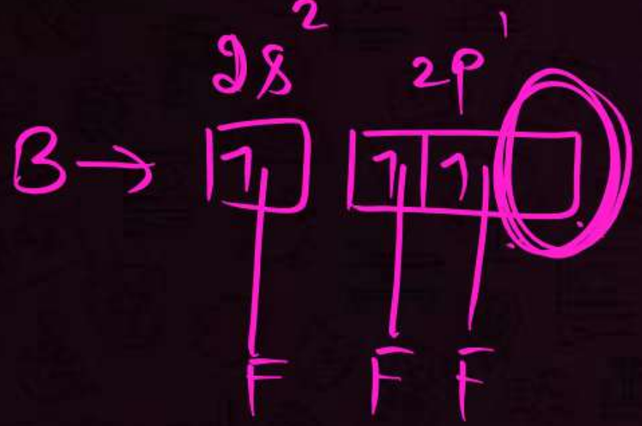
(B) ~~NCl_5~~

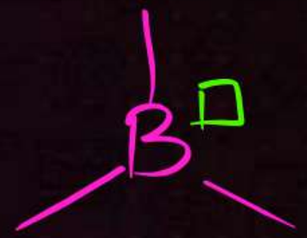
(C) PCl_5

(D) $NH_2^{\ominus}$

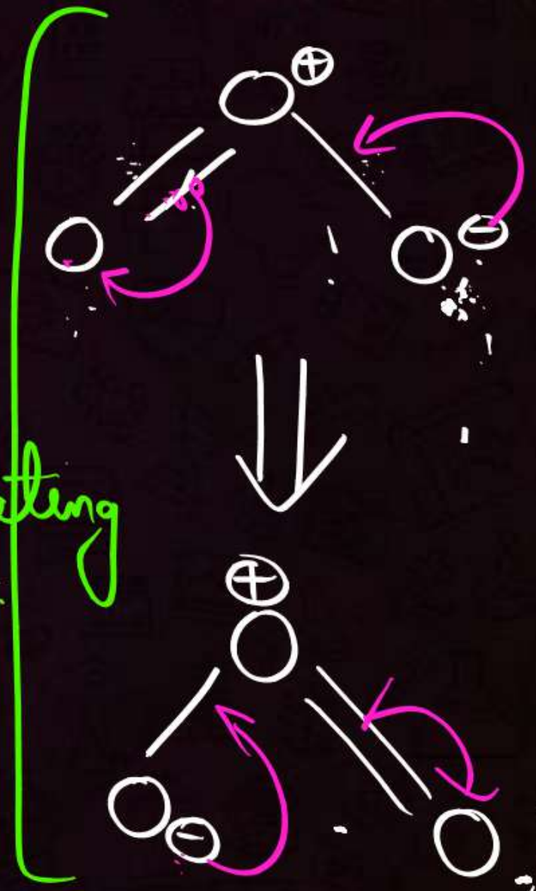
BF_3





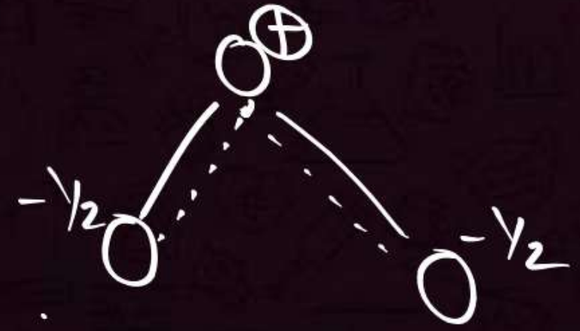


O_3

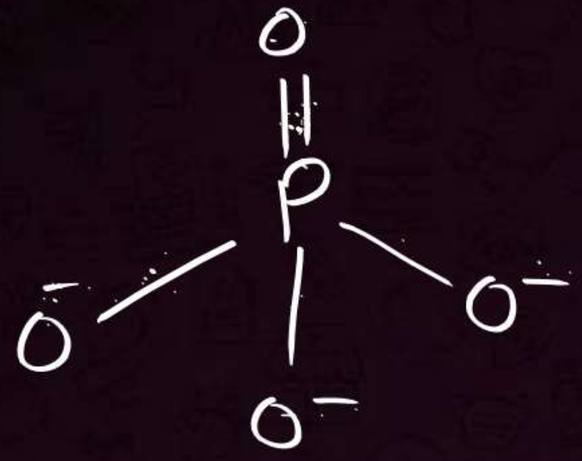


Avg. B.O. $= 1 + \frac{\pi}{\sigma} \Rightarrow 1 + \frac{1}{2} = 1.5$

Avg. F.C. =



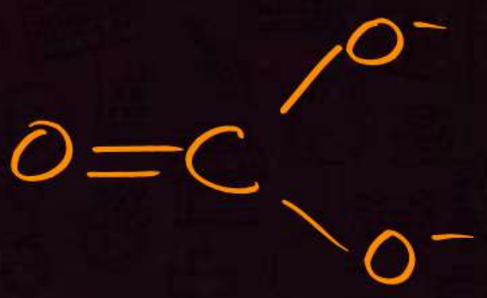
PO_4^{-3}

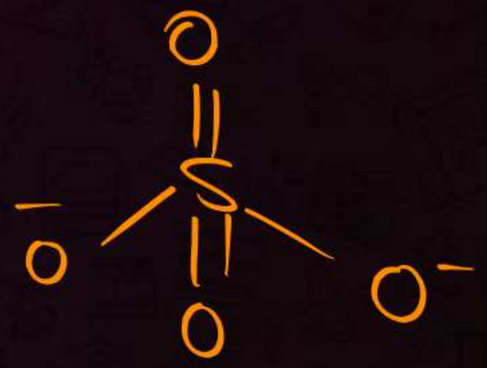


Avg B.O. = $1 + \frac{1}{4} = 1.25$

Avg F.C. = $-3/4$

H.W.

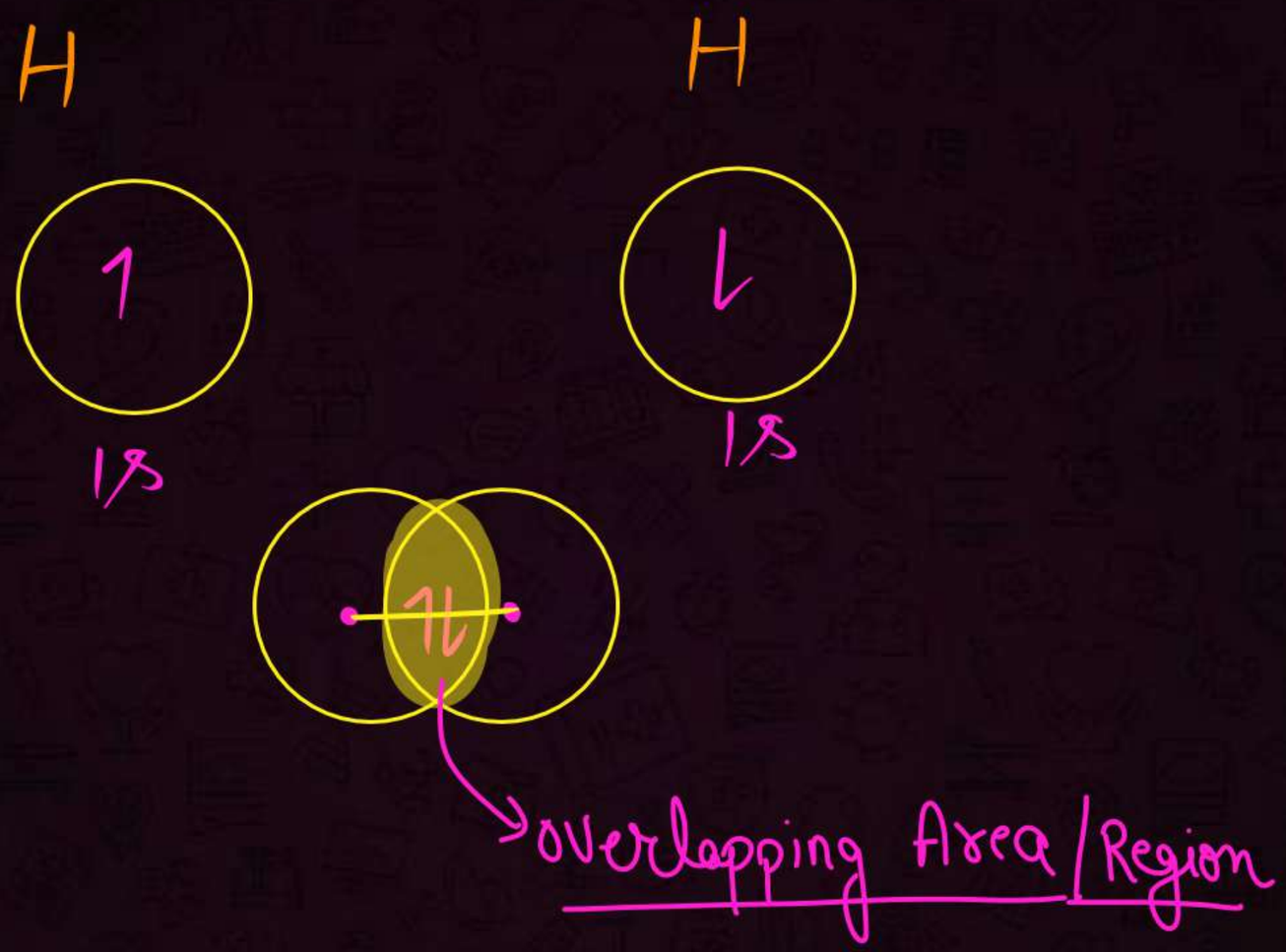




Valence Bond Theory (VBT) (overlapping Concept)

H — H

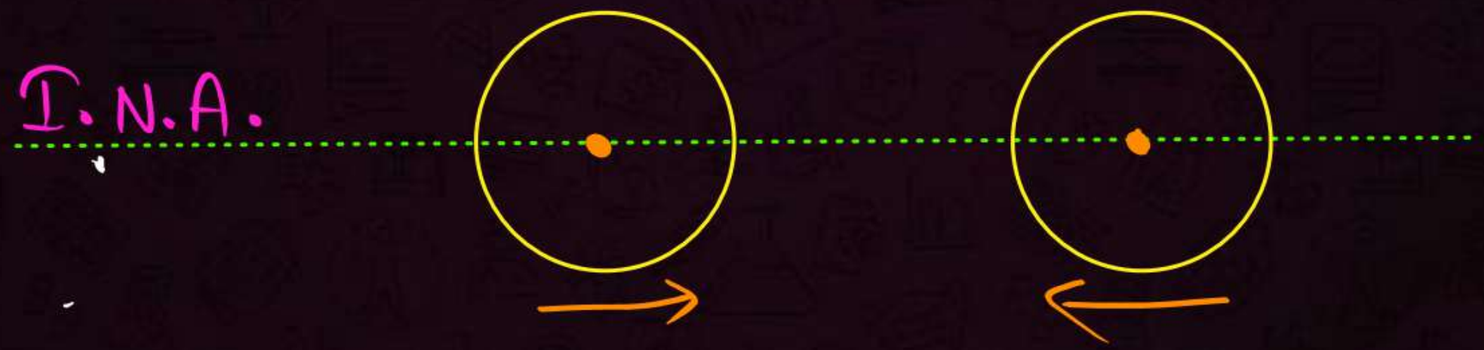
Acc. to V.B.T. a covalent Bond is formed when two half filled Atomic Orbitals (A.O.) having e^- with opposite spin partially combine (overlap) with each other.


H
H
1s
1s
overlapping Area / Region

Types of Covalent Bond

Based on overlapping ⇒

1. σ - Bond ✓
2. π - Bond ✓
3. δ - Bond



NCERT CORNER

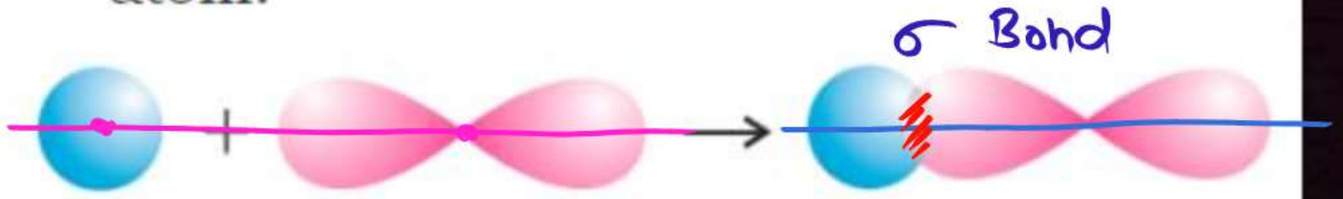
① Axial / Head-on / Head to Head Overlapping :↓

- **s-s overlapping** : In this case, there is overlap of two half filled s-orbitals along the internuclear axis as shown below :

s-orbital + s-orbital → s-s overlapping

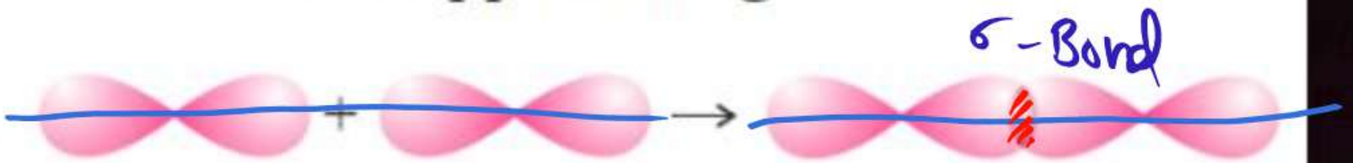
σ - Bond

- ***s-p* overlapping:** This type of overlap occurs between half filled *s*-orbitals of one atom and half filled *p*-orbitals of another atom.



s-orbital *p*-orbital *s-p* overlapping

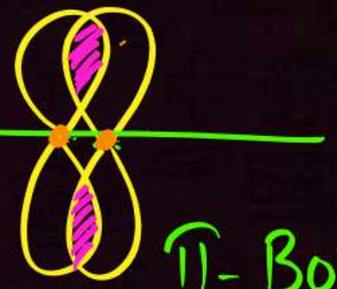
- ***p–p* overlapping** : This type of overlap takes place between half filled *p*-orbitals of the two approaching atoms.



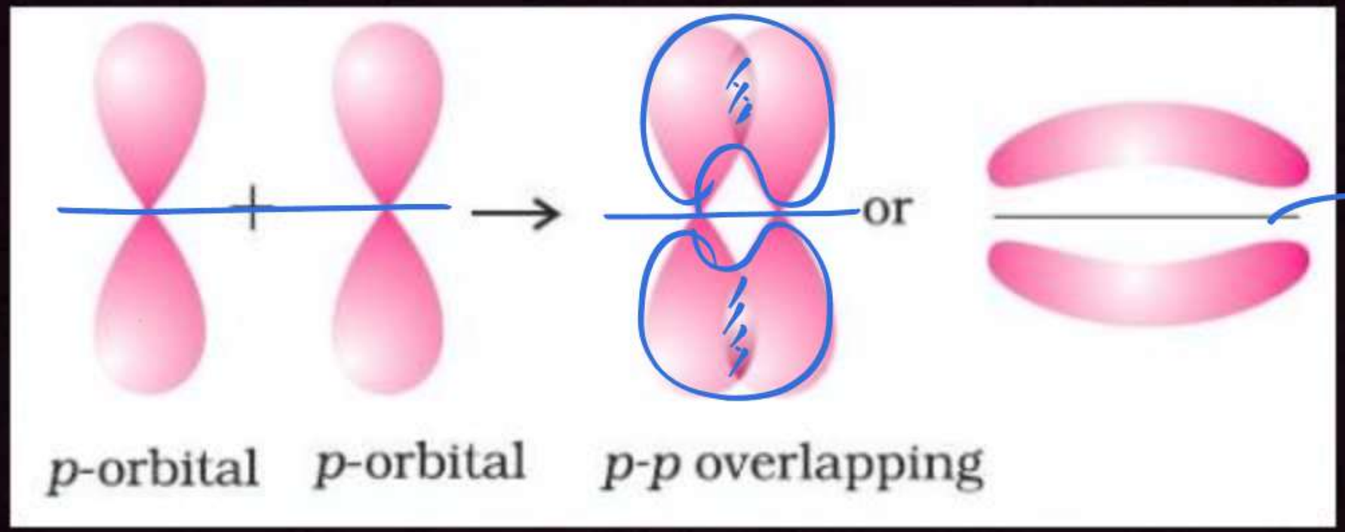
p-orbital *p*-orbital *p-p* overlapping

σ - Bond

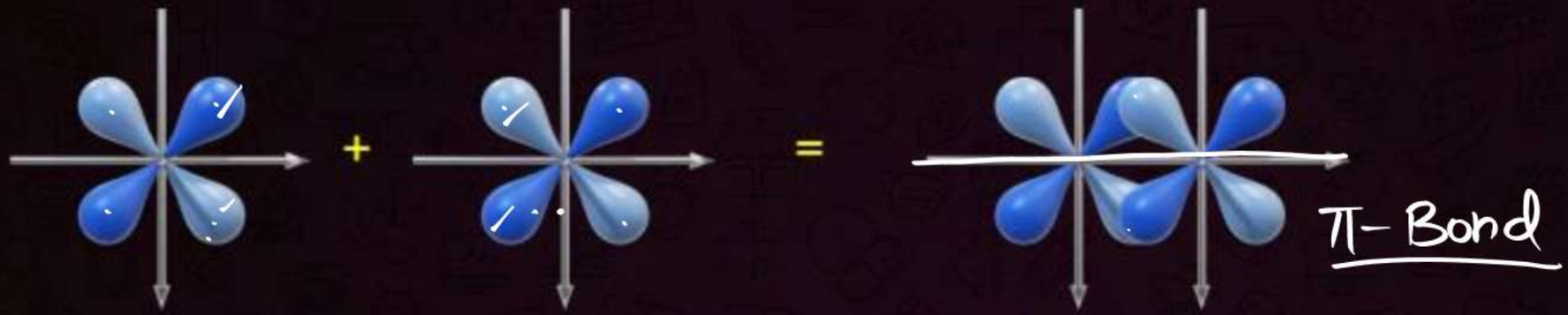
Colateral/
sidways/
side by side

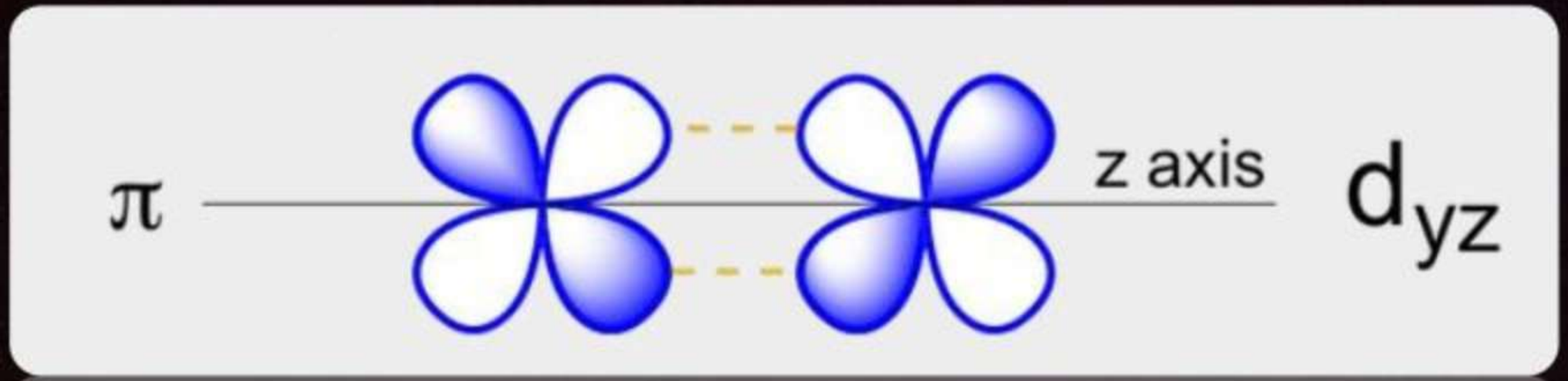


π - Bond

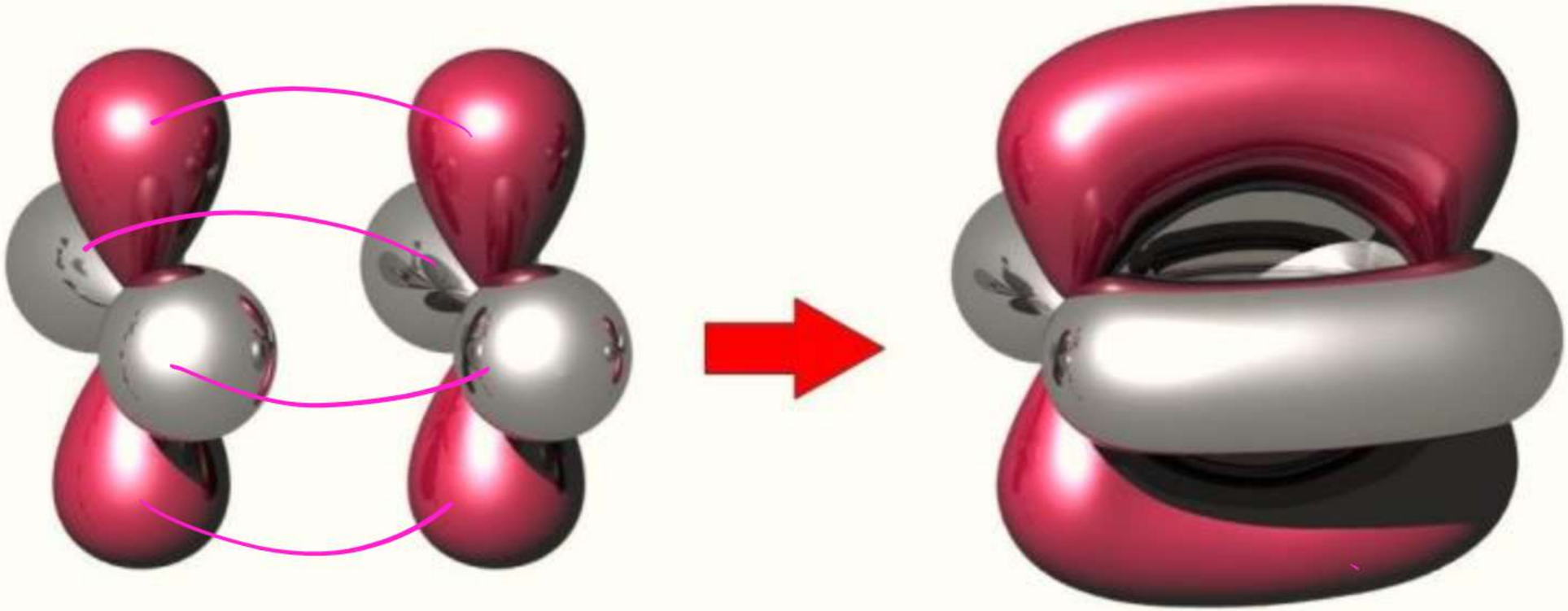

or
Nodal plane
p-orbital
p-orbital
p-p overlapping

Overlapping between d-orbitals :





Delta Bond Type

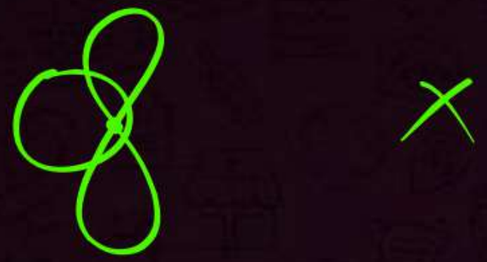


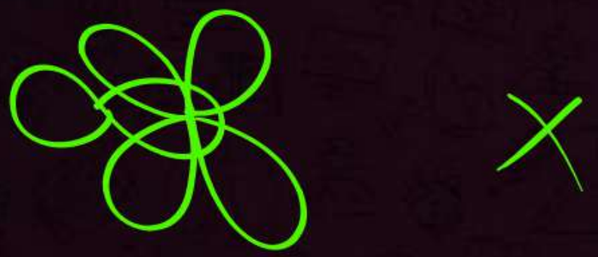
d + d

δd-δd

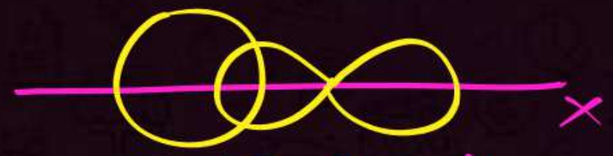
Comparison between σ and p bond

	σ bond		π bond
1	Formed by axial overlapping	1	Formed by side by side overlapping
2	Involves s-s, s-p, p-p (axial) & hybrid orbitals overlapping	2	Involve p-p, p-d & d-d sideways/collateral overlapping
3	Extent of overlapping is stronger σ Stronger	3	Extent of overlapping is less so weaker π- weaker
4	Free rotation around σ bond is possible	4	Free rotation around π bond is not possible
5	Hybridized or unhybridized orbital forms σ bond	5	Hybridized orbital never forms π bond
6	Independent existence of σ-bond.	6	No independent existence.





$s + p_x$ + X-axis $\Rightarrow$



σ-Bond

* tlitricks *

$s + p_z = z \Rightarrow \sigma$

$s + p_x = x \Rightarrow \sigma$

$s + p_x = y \Rightarrow \times$

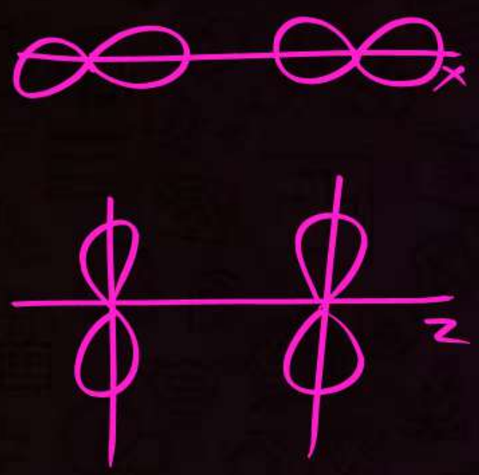
$s + p_z = x \Rightarrow \times$

$s + p_x = z \Rightarrow \times$

$s + p_y = ? = \sigma$

Q. $s + p_z$ orbital form π - Bond on which Axis ?

1. x - axis
2. y - axis
3. z - axis
4. None ✓



$P_x + P_x = \textcircled{x} = \sigma$

$P_y + P_y = y = \sigma$

$P_x + y_x = z = \pi$

$P_x + P_x = y = \pi$

$P_y + P_y = z = \pi$

$P_z + P_z = z = \sigma$

$P_z + P_z = x = \pi$

$P_z + P_z = y = \pi$

$P_x + P_y = x = X$

$P_y + P_z = y = X$

$P_z + P_x = x = X$

$P_z + P_x = z = X$

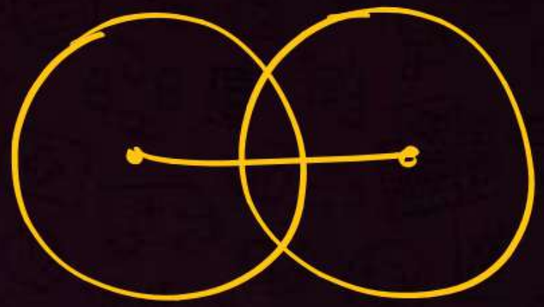
Strength of Covalent Bond

Strength of Covalent Bond $\propto$ Extent of overlapping

Extent of overlapping Depends on

1. size of orbital
2. Directional Nature

① size :→



Bond Strength $\propto \frac{1}{\text{size of orbital}}$

Ex: $1s-1s > 2s-2s > 3s-3s$

$1s-2p > 2s-2p > 3s-2p$

n → change

⇒ use size factor

(II) Directional Nature :-

$s < p < d < f$

s → non directional orbital

Extent of overlapping $\propto$ Directional Nature

NEET

Ex: $2s-2s < 2s-2p < 2p-2p$

n → same

⇒ use dir^{n}. Nature

Ex: $\underbrace{1s-1s}_{n=1} > \underbrace{2s-2s < 2s-2p < 2p-2p}_{n=2} > \underbrace{3s-3s < 3s-3p}_{n=3}$

$\Rightarrow$ $1s-1s > 2p-2p > 2s-2p > 2s-2s > 3s-3p > 3s-3s$

Ex: $3p_\pi - 3p_\pi < 3p_\pi - 3d_\pi < 3d_\pi - 3d_\pi$

Concept of Covalency

→ No. of covalent Bond

covalency

$NCl_3 \longrightarrow 3$

$NH_4^{\oplus} \longrightarrow 4$

$PCl_5 \longrightarrow 5$

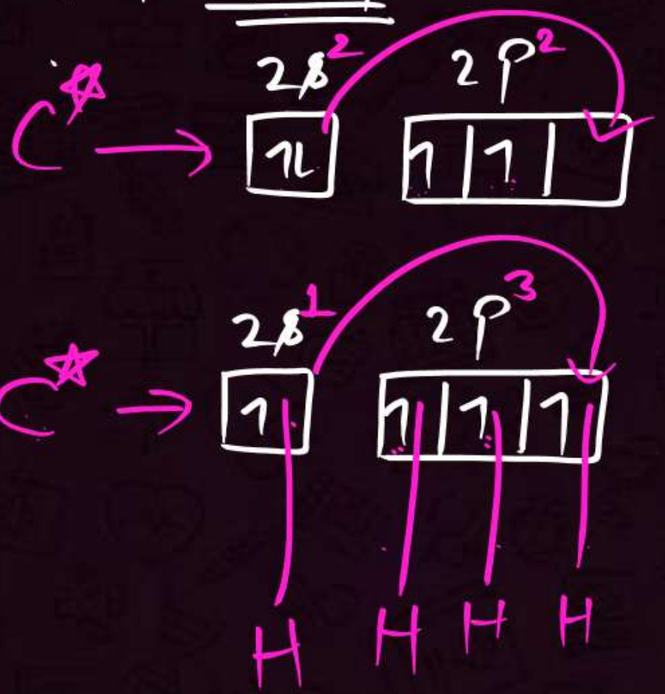
$SF_6 \longrightarrow 6$

Maximum Covalency

2nd period Elements = 4

CH_4 →

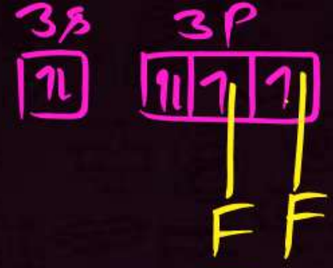
C → $1s^2 2s^2 2p^2$



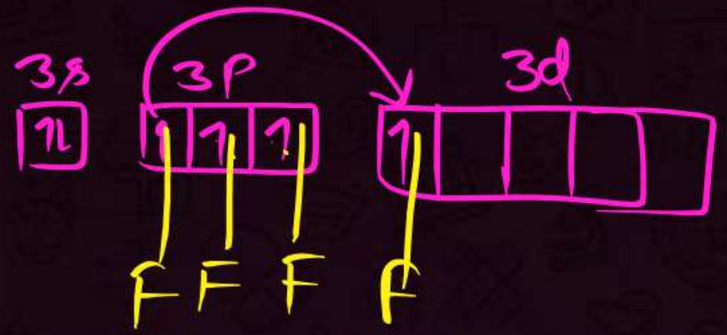
SF_2

~~SF3~~

$S \rightarrow$ (G.S.)

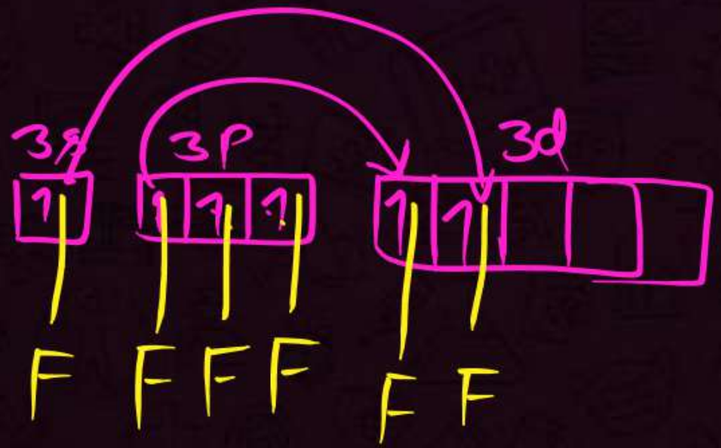


- SF_4 (1st E.S.)



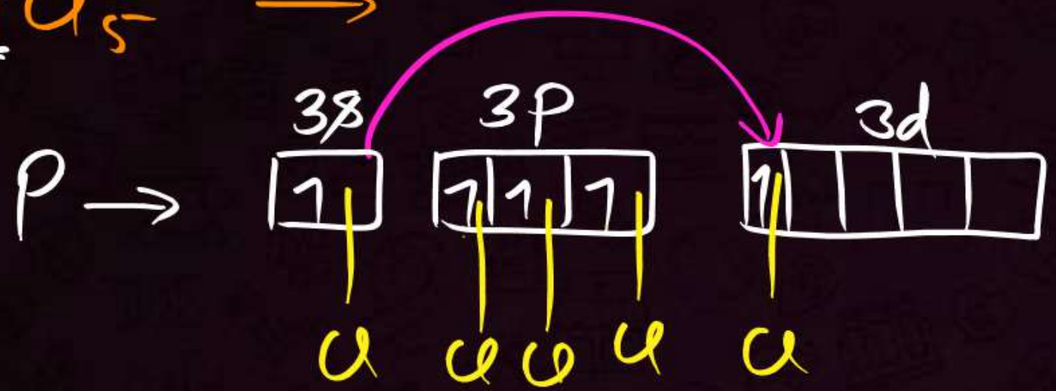
~~SF5~~

- SF_6 (2nd E.S.)



~~JF7~~

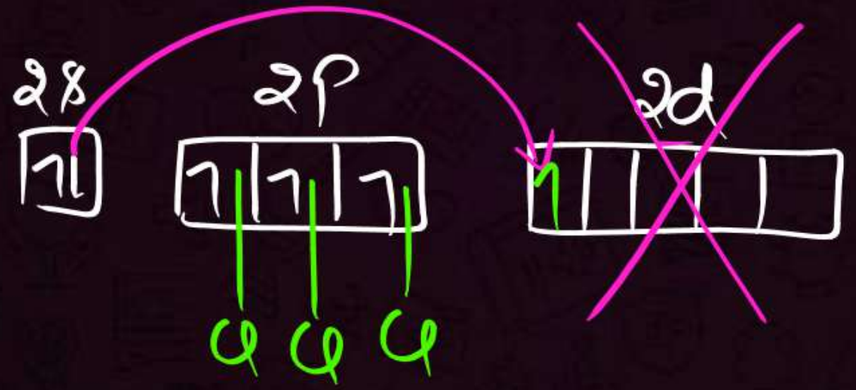
PCl_5 ⟶



PCl_3 PCl_5

~~PCl_4~~

NCl_5 ⟶



No vacant d-orbitals

ClF

ClF_3

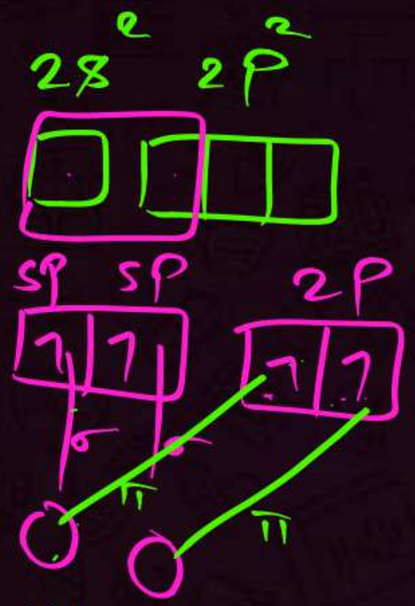
ClF_5

IF_7

H.W.

CO_2

C →



$2p_\pi - 2p_\pi$

100% sure

Hybridisation (Pauling)

Mixing

CH_4

C ↪ 2s | 2p

1 | 1 1 1

sp^3 sp^3 sp^3 sp^3

1 1 1 1

$S\text{-}sp^3$

H H H H

1s

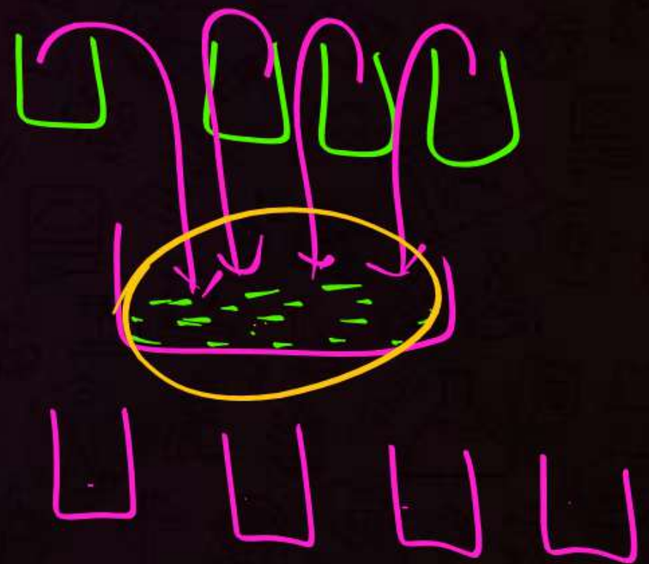
⇒ All overlapping same

→ B.L. & B.S. same.

Teaffy

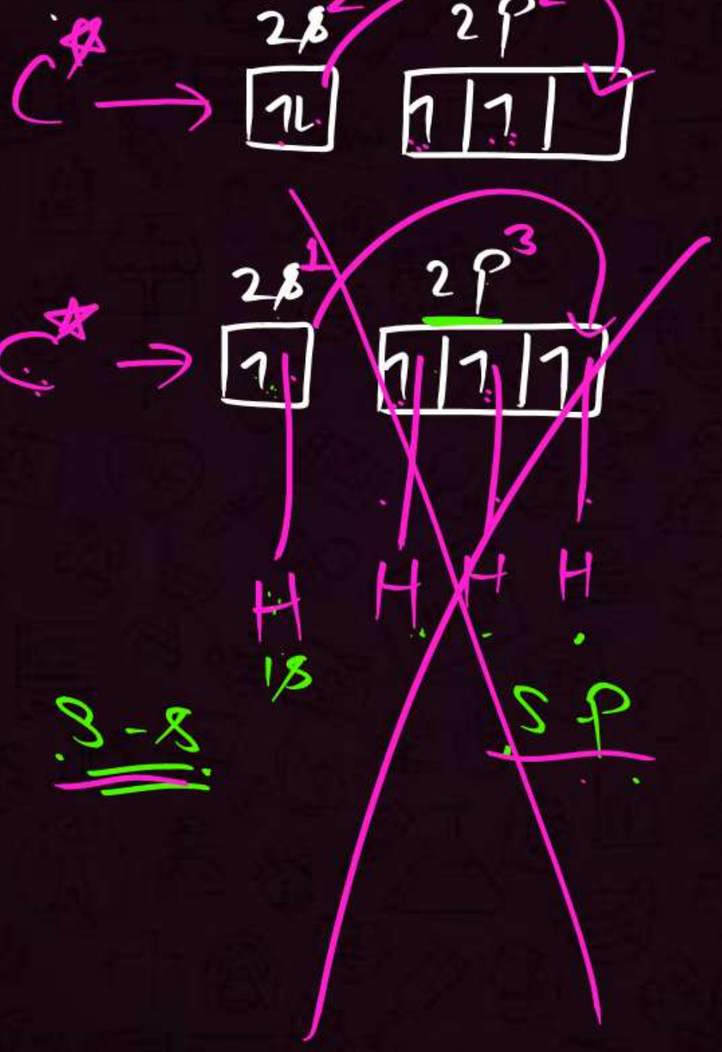
1 1 1 1

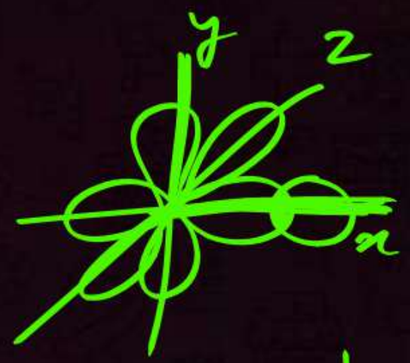
C T T T T



CH_4 →

C → $1s^2 2s^2 2p^2$





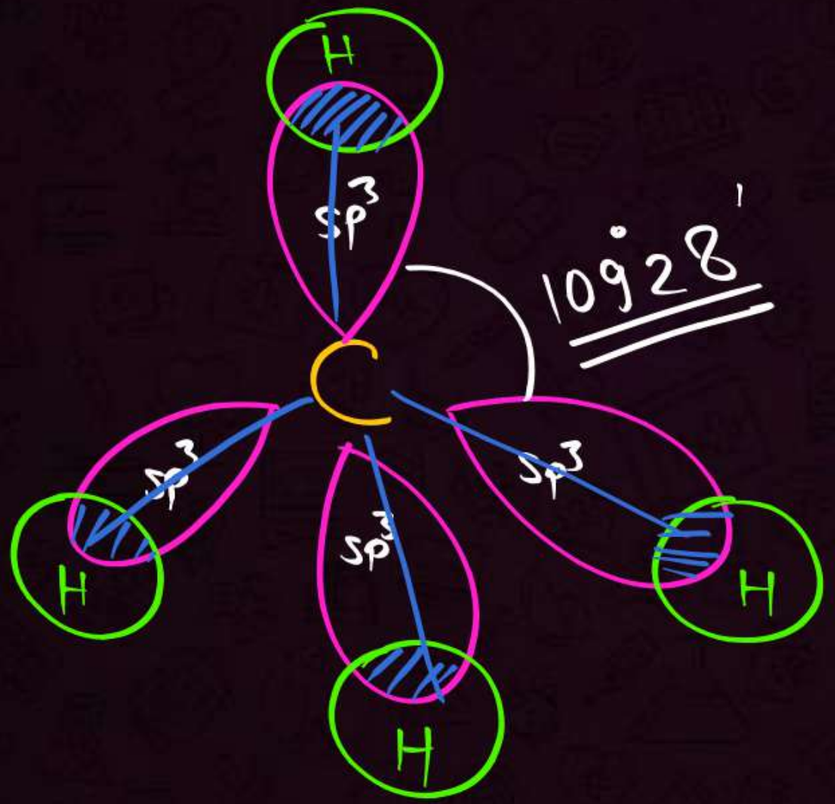
All B.L are same

B.S same

$109^\circ 28'$

Geometry?

$s + p_x + p_y + p_z \Rightarrow 4\underline{sp^3}$



shape | Geometry = Tetrahedural

Salient features of Hybridization :

109°28′ or 109.5°

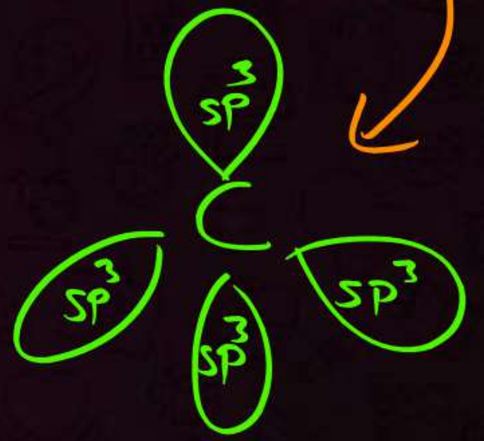
1. The number of hybrid orbitals is equal to the number of the atomic orbitals that get hybridized.

2. The hybridized orbitals are ~~always~~ equivalent in energy and shape.

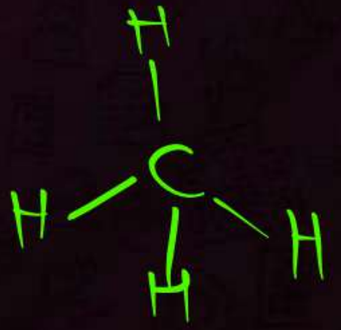
3. The hybrid orbitals are more effective in forming stable bonds than the pure atomic orbitals.

4. These hybrid orbitals are directed in space in some preferred direction to have minimum repulsion between electron pairs and thus a stable arrangement. Therefore, the type of hybridization indicates the geometry of the molecules.

* Arrangement of hybrid orbitals ⇒ electronic Geometry



* Shape of Molecule | Geometry | Molecular Geometry ⇒ Arrangement of bp. only



Calculation of Hybridization = $\underbrace{\sigma + lp}_{n}$

n → Steric No. (no. of Hybrid Orbitals)

Steric No.		Hybridization
2	=	sp
3	=	sp^2
4	=	sp^3
5	=	sp^3d
6	=	sp^3d^2
7	=	sp^3d^3

Calculation of Hybridisation

1. sp hybridisation

$BeCl_2$

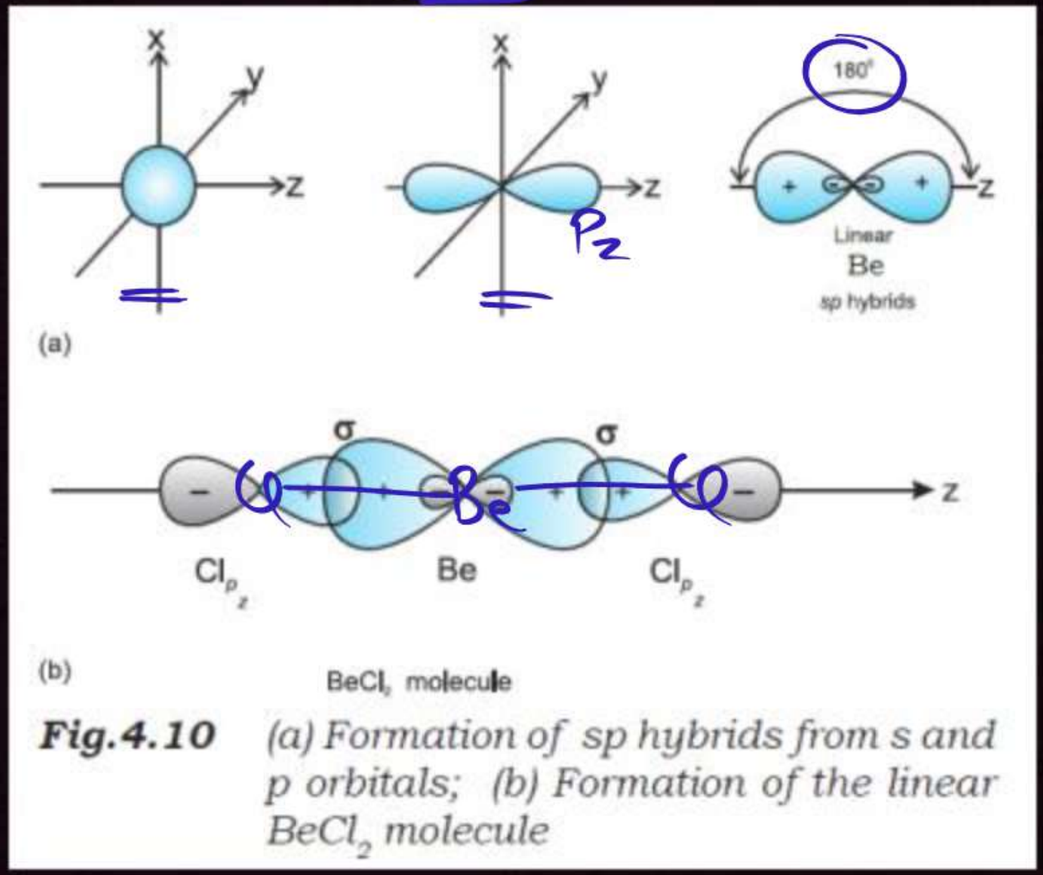


Fig.4.10 *(a) Formation of sp hybrids from s and p orbitals; (b) Formation of the linear $BeCl_2$ molecule*

2. sp^2 hybridisation

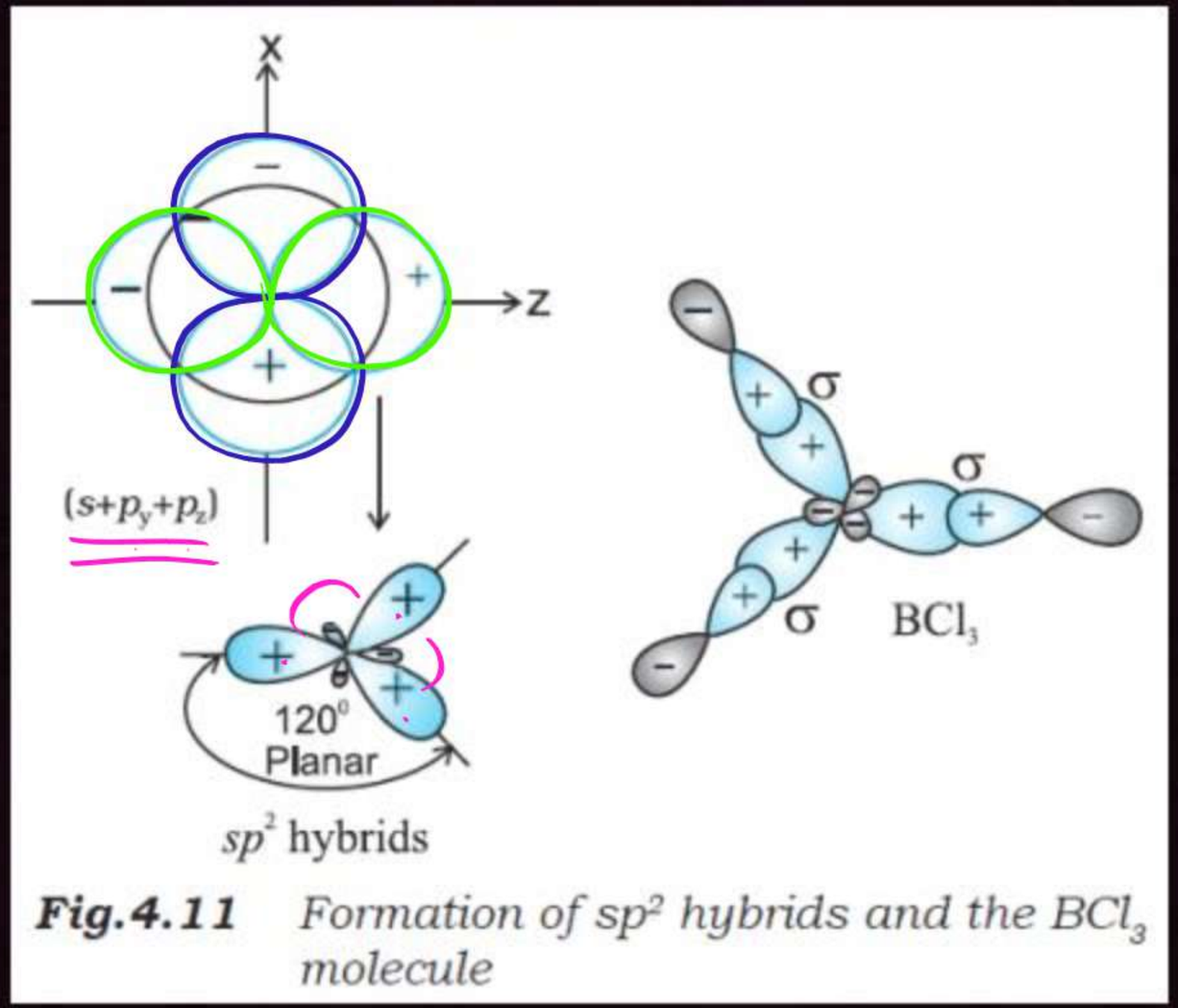


Fig.4.11 *Formation of sp^2 hybrids and the BCl_3 molecule*

3. sp^3 hybridisation

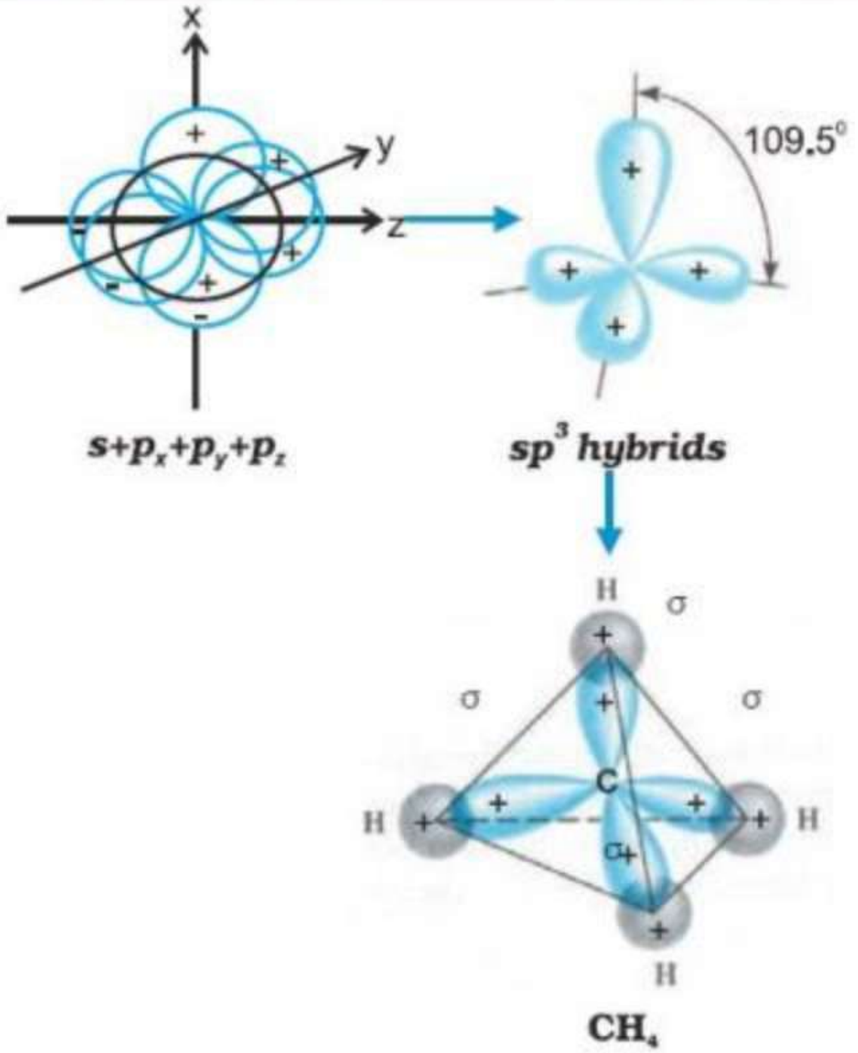
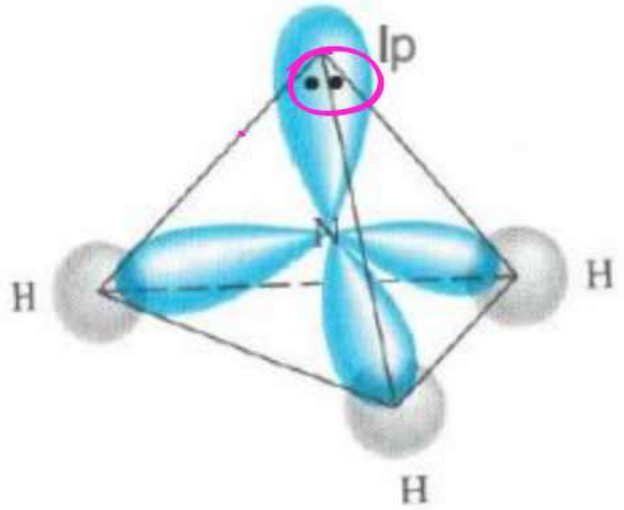


Fig.4.12 *Formation of sp^3 hybrids by the combination of s, p_x, p_y and p_z atomic orbitals of carbon and the formation of CH_4 molecule*



Ammonia, NH_3

Fig.4.13 *Formation of NH_3 molecule*

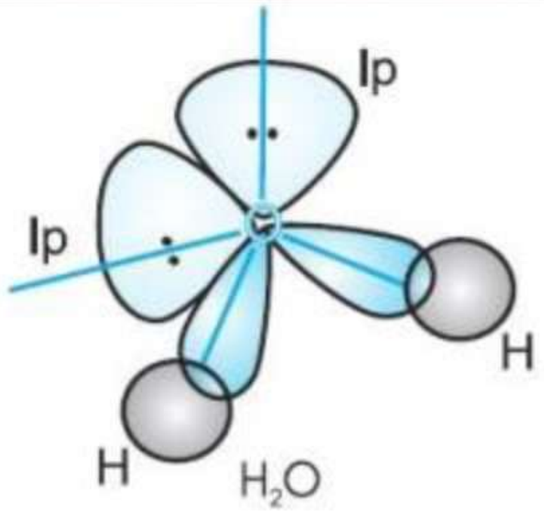
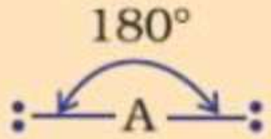
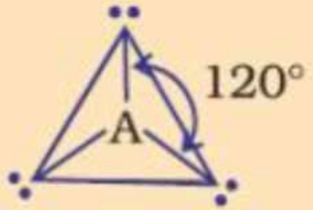
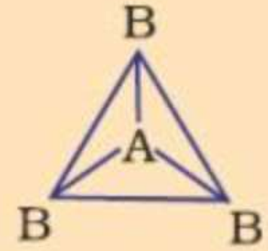
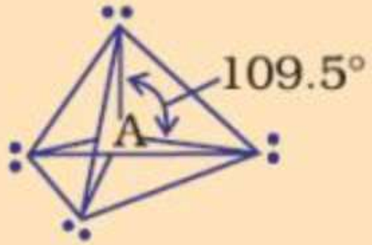
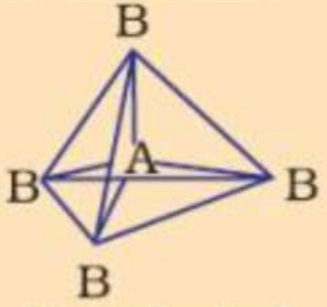
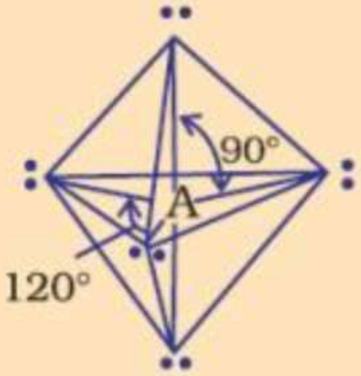
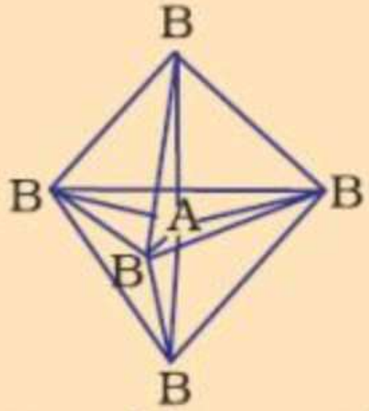
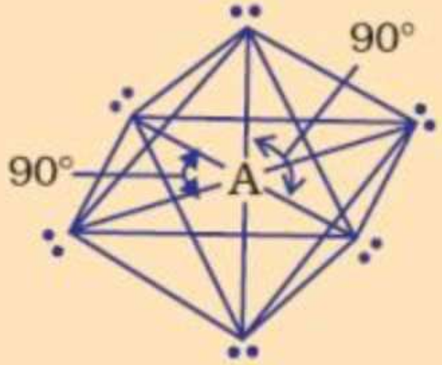
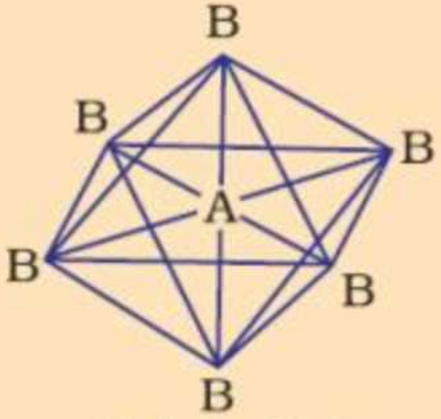


Fig.4.14 *Formation of H_2O molecule*

NCERT CORNER [Page No:- 114]

Number of electron pairs	Arrangement of electron pairs	Molecular geometry	Examples
2	 Linear	B—A—B Linear	$BeCl_2$, $HgCl_2$
3	 Trigonal planar	 Trigonal planar	BF_3

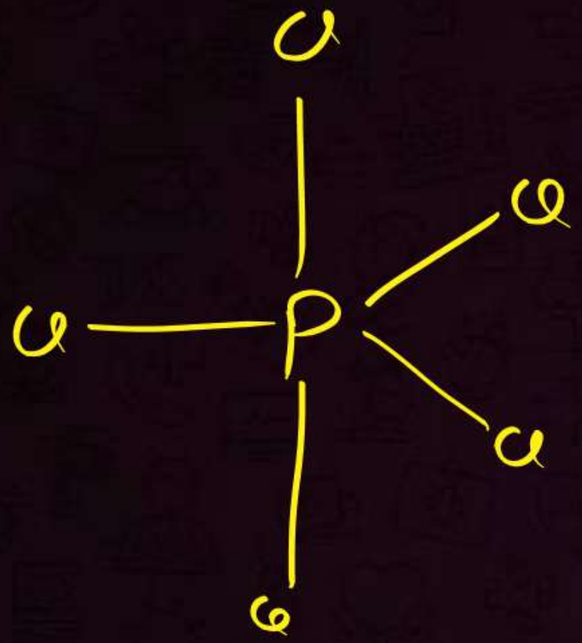
4	 Tetrahedral	 Tetrahedral	CH_4, NH_4^+
5	 Trigonal bipyramidal	 Trigonal bipyramidal	PCl_5
6	 Octahedral	 Octahedral	SF_6

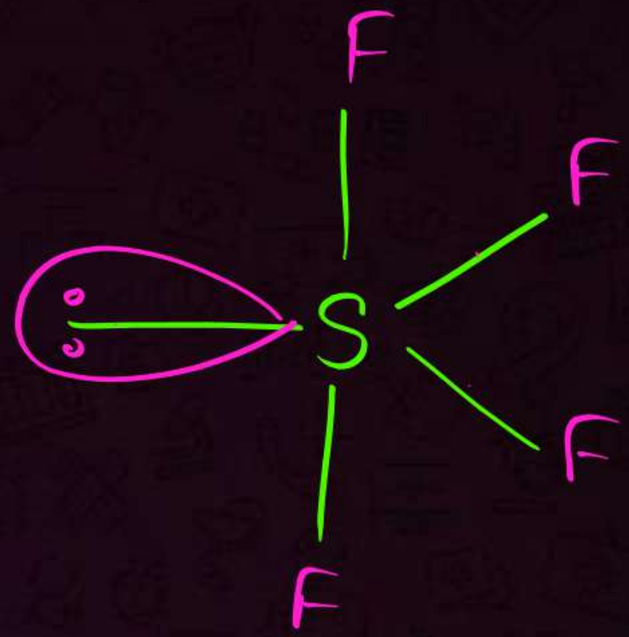
sp^3d ⇒

		e⁻ geometry
$5\sigma + 0lp$	=	T.B.P.
$4\sigma + 1lp$	=	T.B.P.
$3\sigma + 2lp$	=	T.B.P.
$2\sigma + 3lp$	=	T.B.P.

Geometry	
T.B.P.	PCl_5
See-saw	SF_4
Bent-T	ClF_3
linear	XeF_2

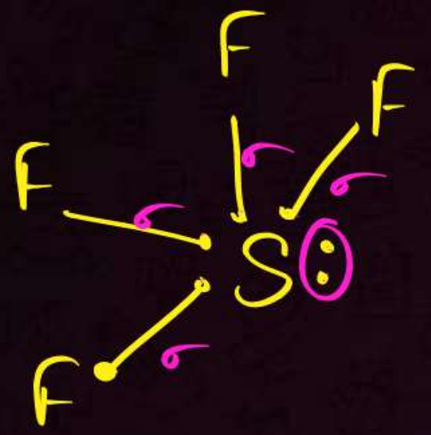
lp ⇒ Equatorial Position



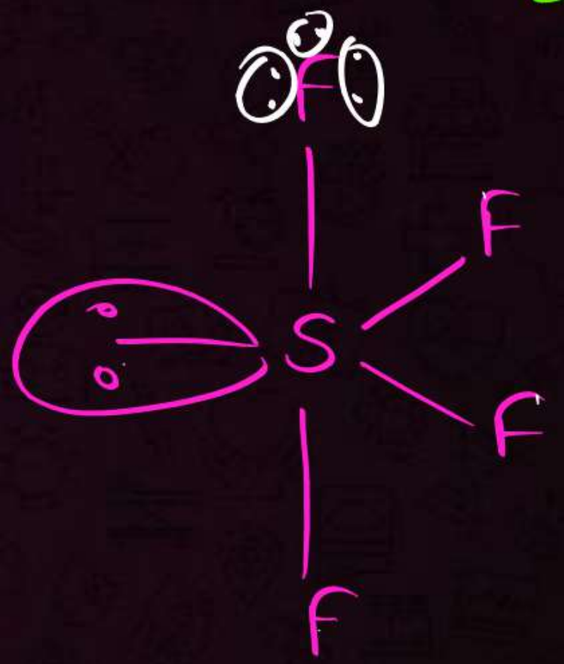


- <u>See-saw</u>
- distorted Tetrahedural
- folded Square

SF_4



$4\sigma + 1lp = 5, sp^3d \rightarrow d_{z^2}$



H, S, O, X

S, O

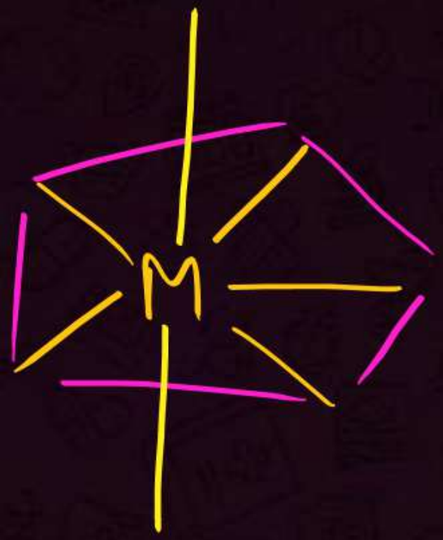
N

$12 + 1 = 13$

		e⁻ geometry	Geometry
sp^3d^2	$6\sigma + 0lp$	octahederal	octahedral
↓	$5\sigma + 1lp$	octahederal	sq. Pyramidal
$d_{x^2-y^2}$			
d_{z^2}	$4\sigma + 2lp$	octahederal	sq. Planar.

sp^3d^3 ⟶ Pentagonal Bipyramidal

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

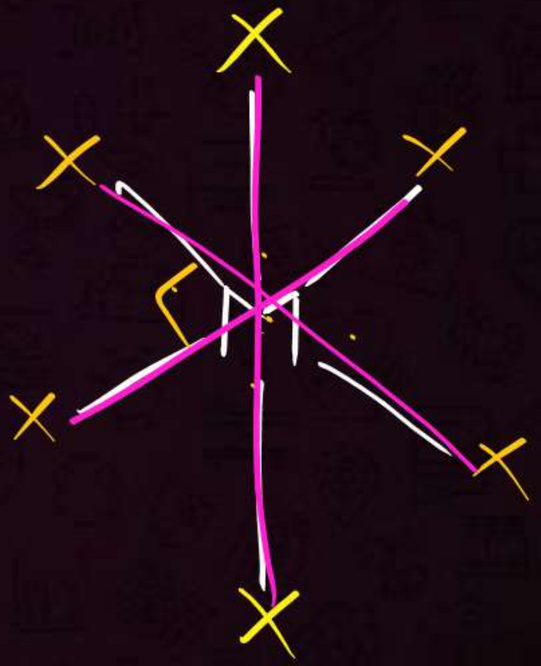


1 lp → distorted octa.

2 lp → Pentagonal planar

QUESTION

In a regular octahedral molecule, MX_6 the number of X–M–X bonds at 180° is :



(A) Two

(B) Six

(C) Four

(D) Three

QUESTION

Specify the coordination geometry around and hybridisation of N and B atoms in a 1:1 complex of BF_3 and NH_3.

(2002, 3M)

(A) N : tetrahedral, sp_3 ; B: tetrahedral, sp^3

(B) N: pyramidal, sp^3; B: pyramidal, sp^3

(C) N: pyramidal, sp^3; B: planar, sp^2

(D) N: pyramidal, sp^3; B: tetrahedral, sp^3

Valence Shell Electron Pair Repulsion Theory

lp → Unshared pair of e⁻

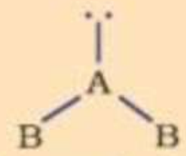
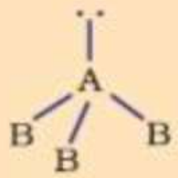
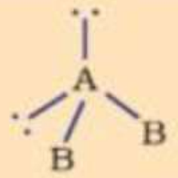
bp → Shared pair of e⁻

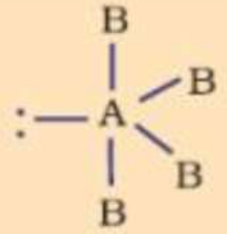
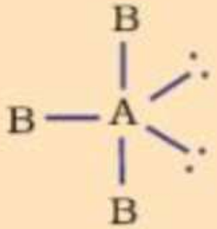
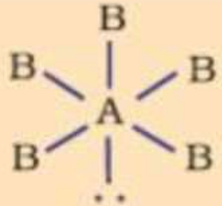
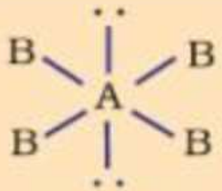
lp-lp > lp-bp > bp-bp

M.B-M.B > M.B-S.B > S.B-S.B

NCERT CORNER [Page No:- 115]

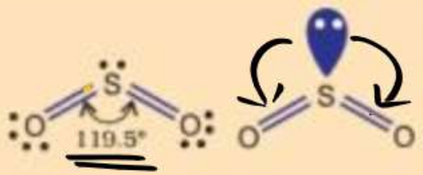
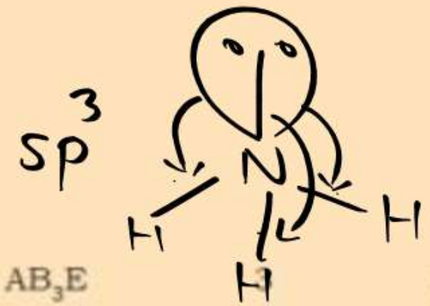
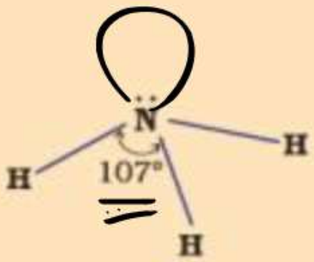
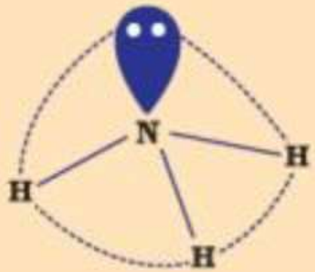
Table 4.7 Shape (geometry) of Some Simple Molecules/Ions with Central Ions having One or More Lone Pairs of Electrons(E).

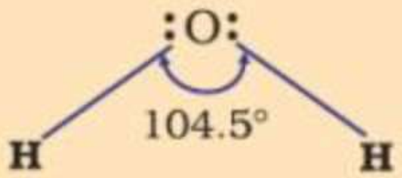
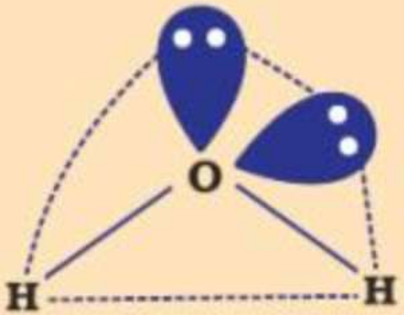
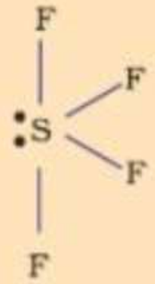
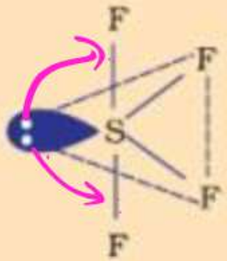
Molecule type	No. of bonding pairs	No. of lone pairs	Arrangement of electron pairs	Shape	Examples
AB_2E	2	1	 Trigonal planer	Bent	SO^2O_3
AB_3E	3	1	 Tetrahedral	Trigonal pyramidal	NH_3
AB_3E_2	2	2	 Tetrahedral	Bent	H_2O

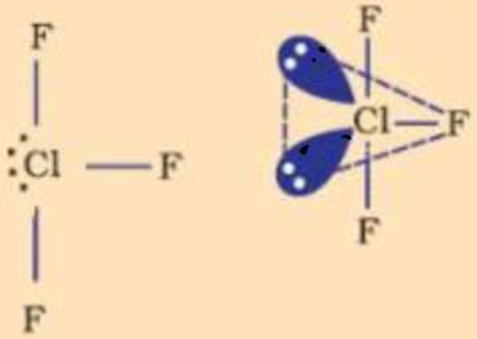
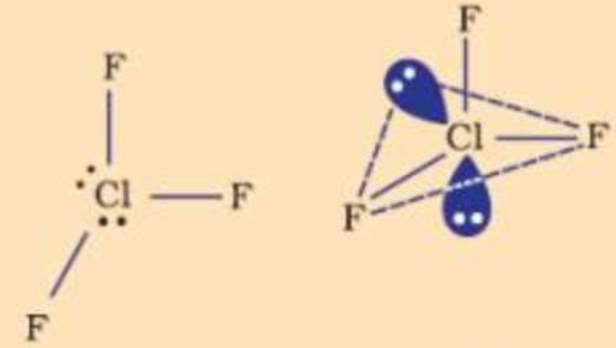
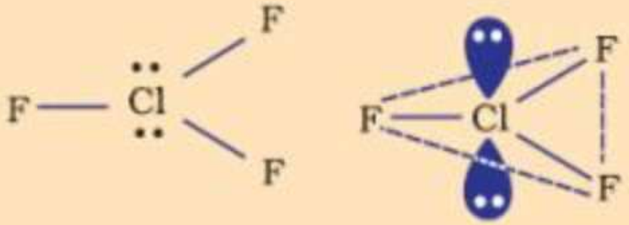
AB_4E	4	1	 Trigonal bi-pyramidal	See saw	SF_4
AB_3E_2	3	2	 Trigonal bi-pyramidal	T-shape	ClF_3
AB_5E	5	1	 Octahedral	Square pyramid	BrF_5
AB_4E_2	4	2	 Octahedral	Square planer	XeF_4

NCERT CORNER [Page No:- 116-117]

Table 4.8 Shapes of Molecules containing Bond Pair and Lone Pair

Molecule type	No. of bonding pairs	No. of lone pairs	Arrangement of electrons	Shape	Reason for the shape acquired
AB_2E	4	1	 	Bent	Theoretically the shape should have been triangular planar but actually it is found to be bent or v-shaped. The reason being the lone pair-bond pair repulsion is much more as compared to the bond pair-bond pair repulsion. So the angle is reduced to 119.5° from 120°.
AB_3E	3	1	  	Trigonal pyramidal	Had there been a bp in place of lp the shape would have been tetrahedral but one lone pair is present and due to the repulsion between lp-bp (which is more than bp-bp repulsion) the angle between bond pairs is reduced to 107° from 109.5°.

AB_2E_2	2	2	 	Bent	The shape should have been tetrahedral if there were all bp but two lp are present so the shape is distorted tetrahedral or angular. The reason is lp-lp repulsion is more than lp-bp repulsion which is more than bp-bp repulsion. Thus, the angle is reduced to 104.5° from 109.5°.
AB_4E	4	1	(a)  (b)   (More stable)	See-saw	In (a) the lp is present at axial position so there are three lp—bp repulsions at 90°. In(b) the lp is in an equatorial position, and there are two lp—bp repulsions. Hence, arrangement (b) is more stable. The shape shown in (b) is described as a distorted tetrahedron, a folded square or a see-saw.

Molecule type	No. of bonding pairs	No. of lone pairs	Arrangement of electrons	Shape	Reason for the shape acquired
AB_3E_2	3	2	(a)  (b)  (c) 	T-shape	In (a) the lp are at equatorial position so there are less lp-bp repulsions as compared to others in which the lp are at axial positions. So structure (a) is most stable. (T-shaped).

QUESTION

Total number of lone pair of electron in I_3^- ion is

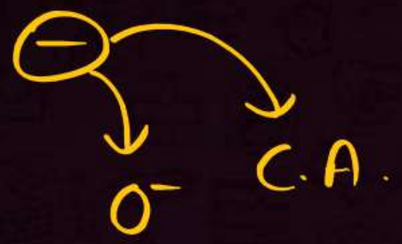
(2018 Main)

A 3

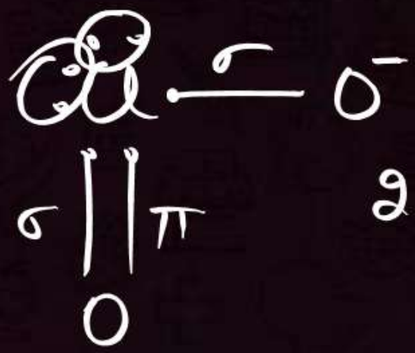
B 6

C 9

D 12



ClO_2^-



$2\sigma + 2lp = 4$

sp^3

Bent

$2\sigma + 3lp = 5, sp^3d$

linear.

$NO_2^\ominus$

sp^2; 1lp

Bent.

QUESTION

Which of the following contains maximum number of lone pairs on the central atom?

(2005, 1M)

3σ, 1lp sp^3

(A) ClO_3^- (8)

(B) XeF_4 (14)

(C) SF_4 (13)

(D) I_3^- (9)

QUESTION

Number of lone pair(s) in $XeOF_4$ is/are

(2004, 1M)

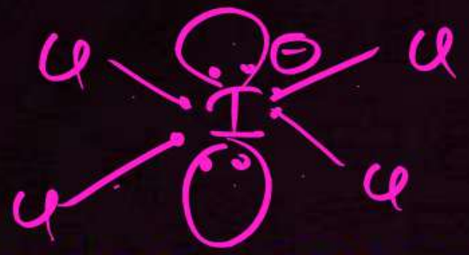
A 0

B 1

C 2

D 3

QUESTION



sp^3d^2

The correct statement about ICl_5 and ICl_4^-, is

Sq. pyr. Sq. planar

(2019 Main, 8 April II)

(A) ICl_5 is square pyramidal and ICl_4^- is tetrahedral

(B) ICl_5 is square pyramidal and ICl_4^- is square planar

(C) Both are isostructural

(D) $IC1_5$ is trigonal bipyramidal and ICl_4^- is tetrahedral

QUESTION

The ion that has sp^3d^2 hybridisation for the central atom, is

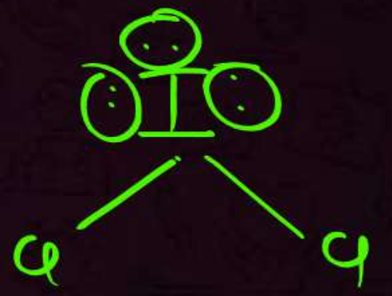
(2019 Main, 8 April)

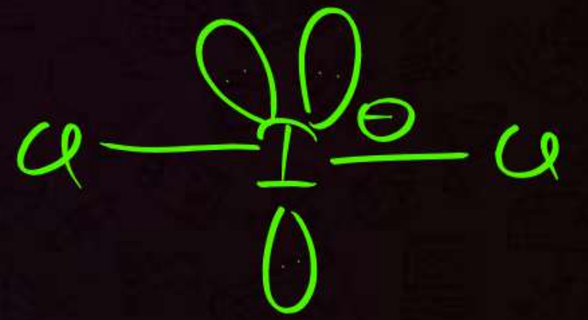
A $[ICl_2]^-$ ⟶ sp^3d

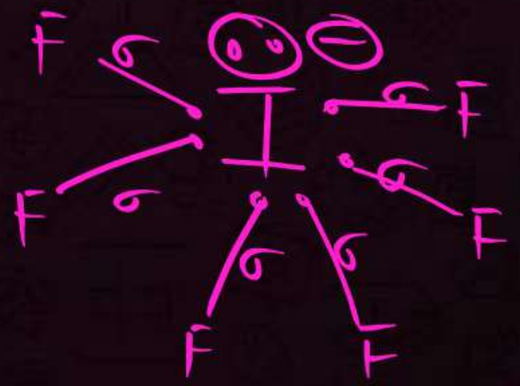
B $[BrF_2]^-$ ⟶ sp^3d

C $[ICl_4]^-$ ⟶ sp^3d^2

D $[IF_6]^-$ ⟶ sp^3d^3







$6\sigma + 1lp = 7$ sp^3d^3

Q IF_6^- d-orbitals?

sp^3d^3

d_{xy}

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

d_{z^2}

QUESTION

Which one of the following molecules is planar? **(1996, 1M)**

(A) NF_3

(B) NCl_3

(C) PH_3

(D) BF_3

Bond Parameters

- Bond Angle
- Bond length
- Bond Strength

Bond Angle

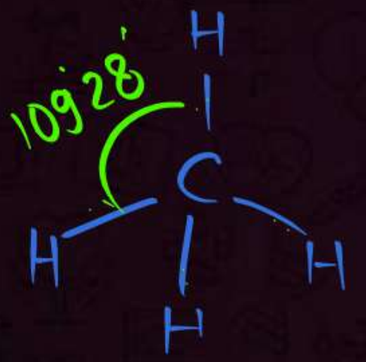
① Hybridization / Geometry :→ (% s character)

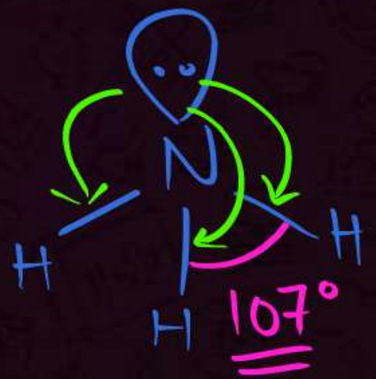
sp	sp^2	sp^3
180°	120°	109°28′
1 s = 50%	33.3%	25%

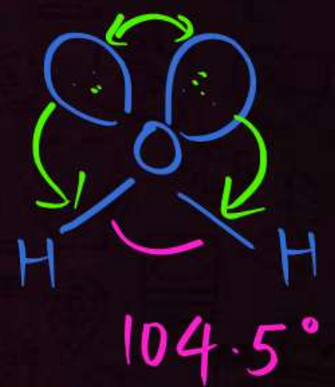
B.A ∝ % s char.

② No. of lp :→

	sp^3	sp^3	sp^3
	CH_4	NH_3	H_2O
no. of lp =	0	1	2



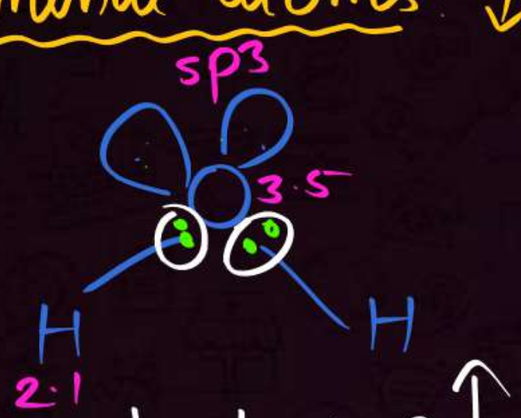




$$B.A \propto \frac{1}{\text{no. of lp. on C.A.}}$$

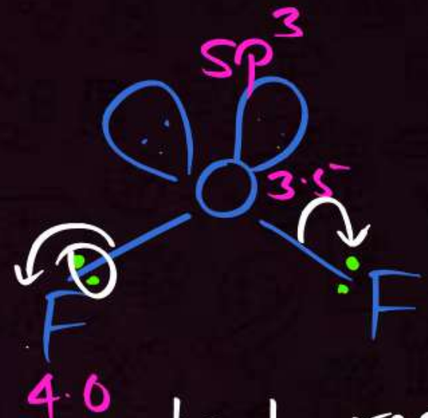
1 lp ↑ B.A ↓ by 2.5°

③ E.N. of Terminal atoms ↘



bp-bp rep ↑

B.A ↑

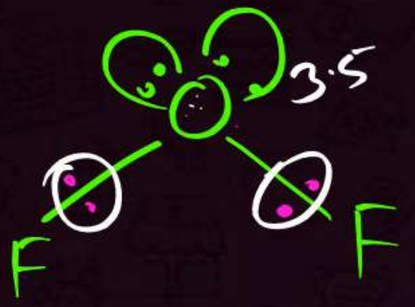


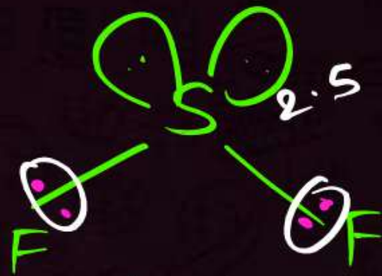
bp-bp rep ↓

B.A ↓

$$B.A \propto \frac{1}{\text{E.N. of Terminal atom}}$$

④ E.N. of Central Atom :-↓





bp-bp rep ↓

B.A ↓

B.A. ∝ E.N. of C.A.

Ex:

$NH_3 > PH_3 > AsH_3 > SbH_3$

$H_2O > H_2S > H_2Se > H_2Te$

$\approx 90^\circ$

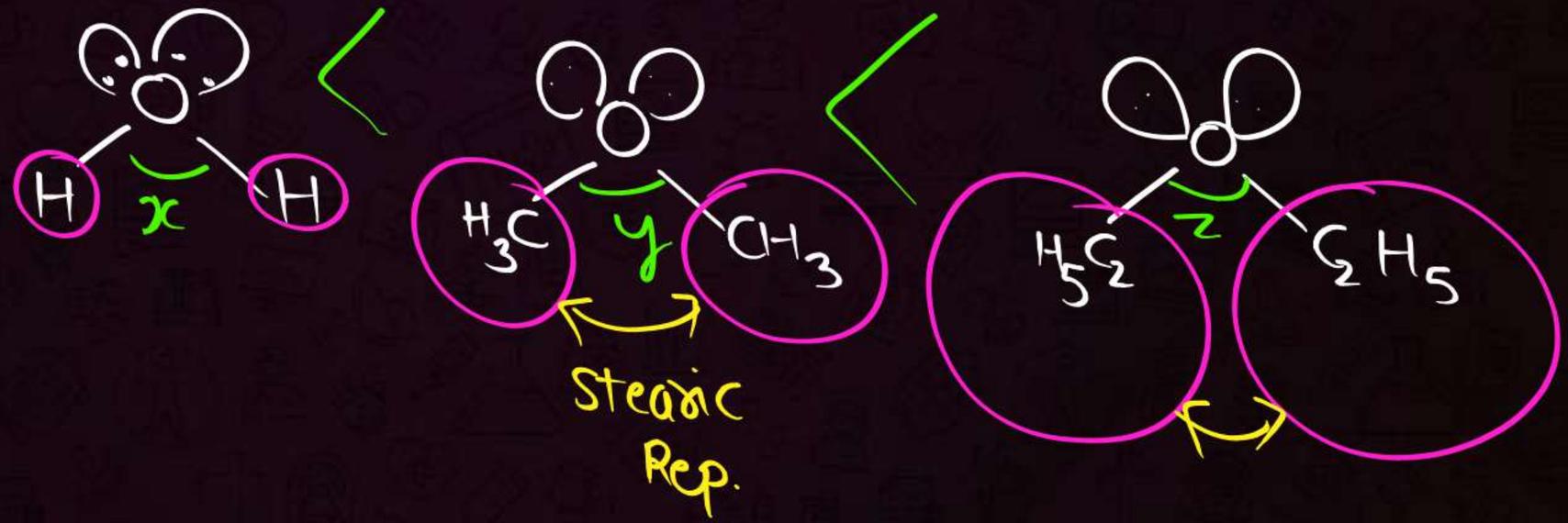
⟹ Drago's Rule

No-hybridization

$NCl_3 > PCl_3 > AsCl_3 > SbCl_3$

sp^3

(5) Size of Terminal atoms :→



E.N.

$H_2O > OF_2$

1 2

Steric Rep

$OCl_2 < OBr_2$

3 4

$OF_2 < H_2O < OCl_2 < OBr_2$

sp^3

2nd period Element

⇓

3rd period or below

or

Alky Group

Q Maximum Bond Angle is Present in ?

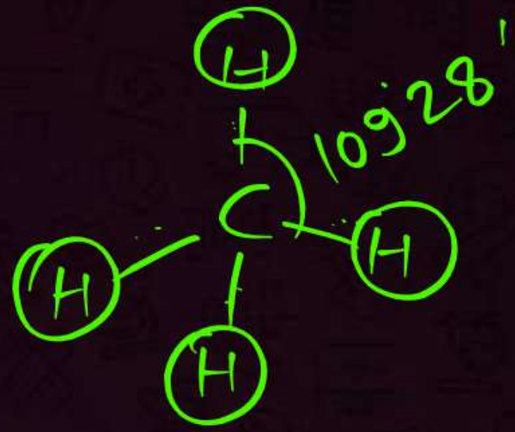
(1) CH_4

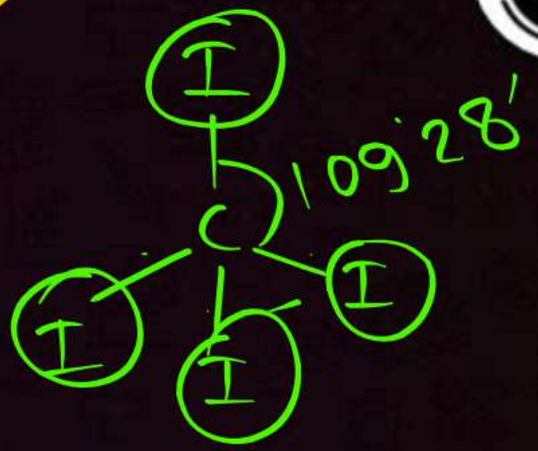
(2) CF_4

(3) CBr_4

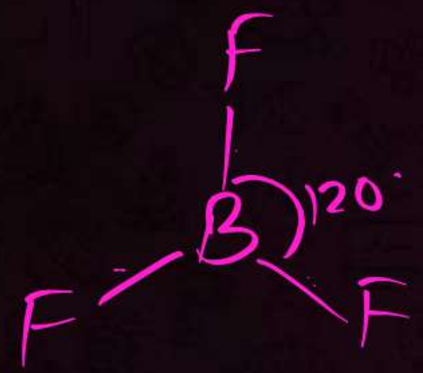
(4) CI_4

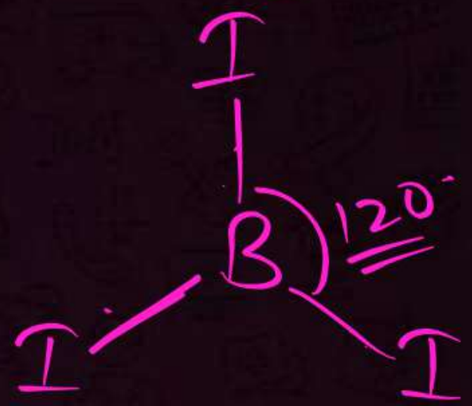
(5) All Equal.





B.A ⇒ BF_3 = BCl_3 = BBr_3 = BI_3





BF_3 (sp^2, 120°) > CCl_4 (sp^3, 109°28′)

Bond Length

① Size :→ B.L. $\propto$ size of atom

$$HF < HCl < HBr < HI$$

② % schar :→

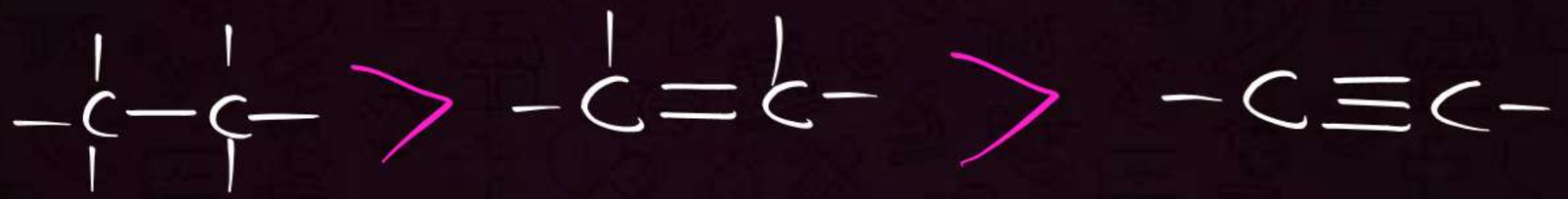
sp	=	50%
sp^2	=	33.3%
sp^3	=	25%

B.L $\propto \frac{1}{\%\,s}$

B.L $\propto$ % pchar.

(3) Bond Order :→

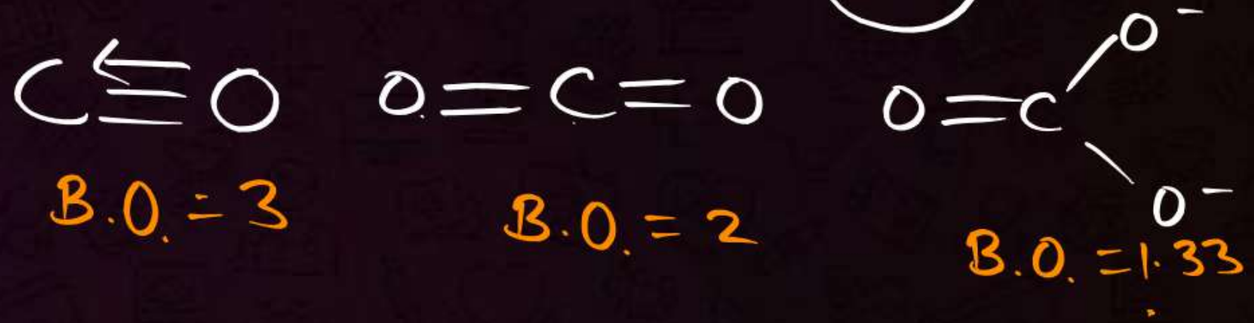
$$B.L \propto \frac{1}{B.O.}$$



QUESTION

The correct order of increasing C – O bond length of CO, CO_3^{2-} < CO_3 is

(1999, 2M)



(A) $CO_3^{2-} < CO_2 < CO$

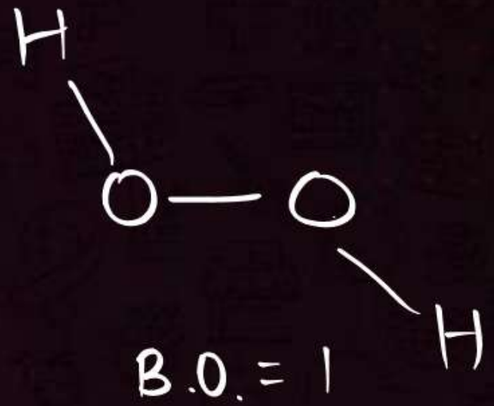
(B) $CO_2 < CO_3^{2-} < CO$

(C) $CO < CO_3^{2-} < CO_2$

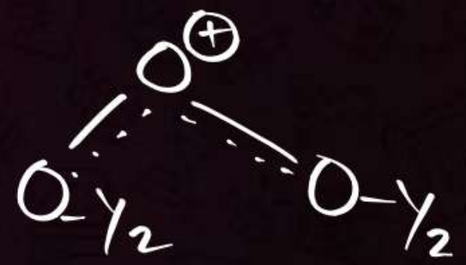
(D) $CO < CO_2 < CO_3^{2-}$

Q B.L order ?

$O_2 < O_3 < H_2O_2$



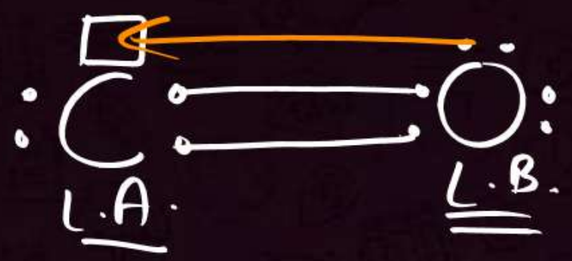
B.O. = 1



B.O. = 1.5

O = O

B.O. = 2


L.A.
L.B.

Bond Strength

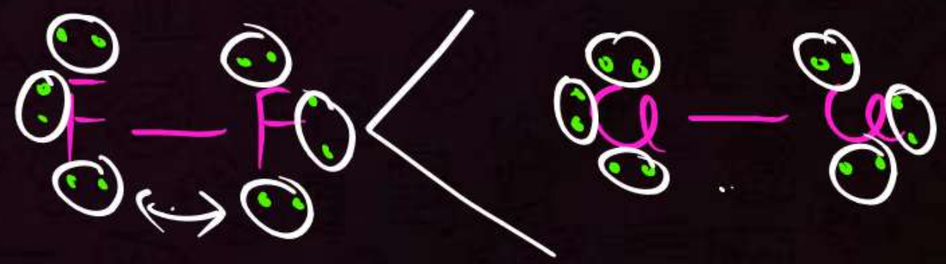
① B.L. :→

$$B.S. \propto \frac{1}{B.L.}$$

② B.O. ⇒

$$B.S. \propto B.O.$$

③ lp-lp Repulsion :→



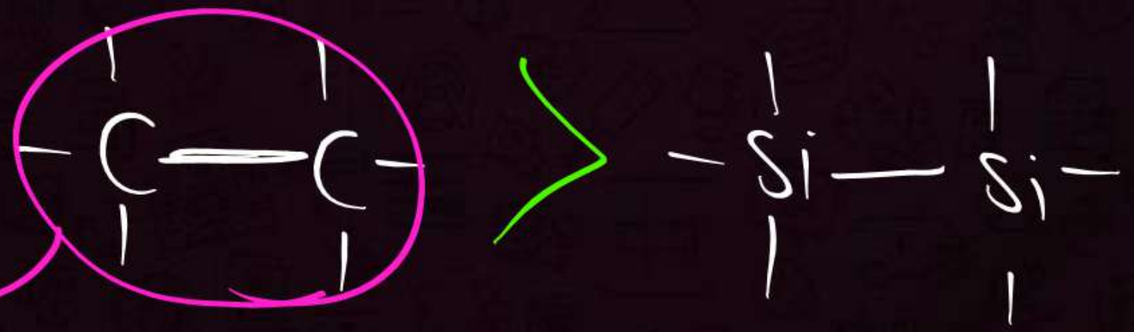
B.S. $\Rightarrow$ F–F < Cl–Cl > Br–Br > I–I

Cl–Cl > Br–Br > F–F > I–I

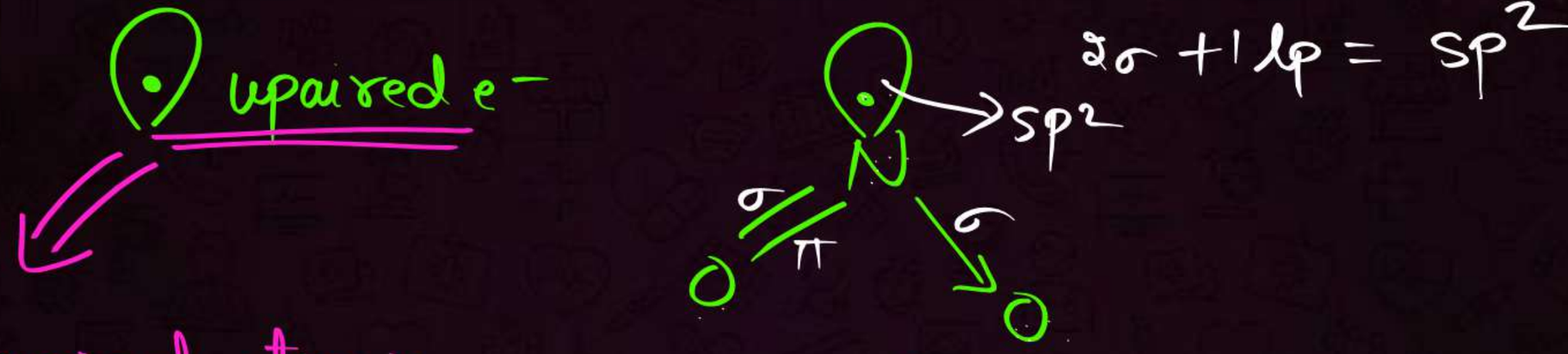
B.S. $\Rightarrow$ $\ddot{N}-\ddot{N}$ < $\ddot{P}-\ddot{P}$

B.S. $\Rightarrow$ $\ddot{O}-\ddot{O}$ < $\ddot{S}-\ddot{S}$

Catenation

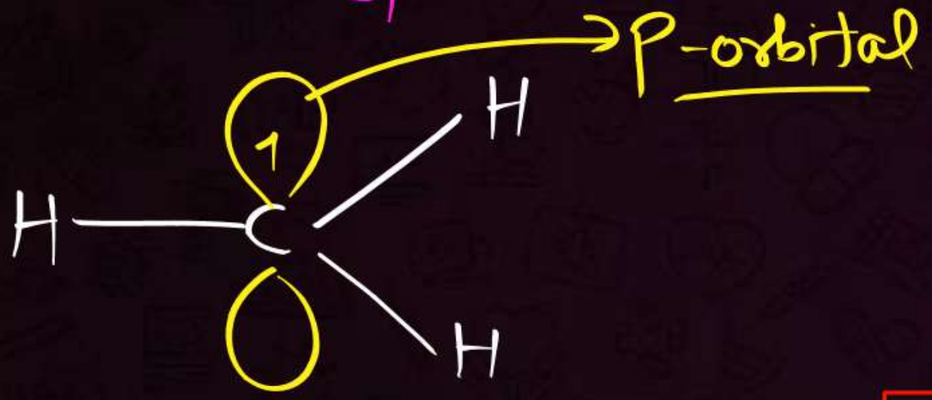
B.S.  >

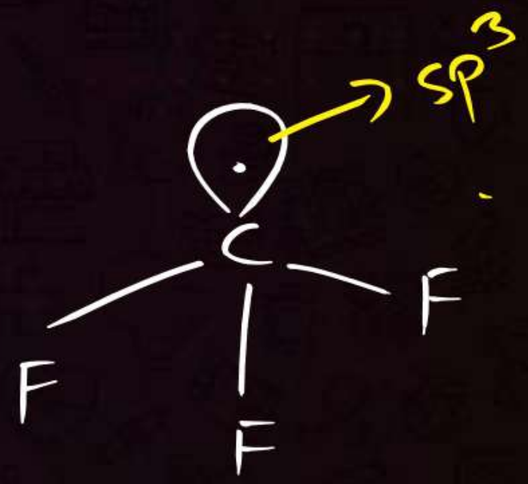
Hybridisation in Odd Electron Species



Terminal atom is more E.N. than C.A. then count the Unpaired e⁻ in hybridization.

& if Terminal atom is less E.N. than C.A. then do not count Unpaired e⁻ in hybridization.





$CO_2 \Rightarrow sp^2$

$CO_3 \Rightarrow sp^3$

Hybridisation in Solid State

$PCl_{5(g)} \rightarrow sp^3d$

$PCl_{5(s)} \longrightarrow \underset{sp^3}{PCl_4^{\oplus}} + \underset{sp^3d^2}{PCl_6^{\ominus}}$

$PBr_{5(s)} \longrightarrow \underset{sp^3}{PBr_4^{\oplus}} + \underline{Br^-}$

$N_2O_3 \longrightarrow \underline{NO^+} + \underset{sp^2}{NO_2^-}$

$N_2O_5 \longrightarrow \underset{sp}{NO_2^+} + \underset{sp^2}{NO_3^-}$

$Cl_2O_6 \longrightarrow \underset{sp^2}{ClO_2^+} + \underset{sp^3}{ClO_4^-}$

$XeF_6 \longrightarrow \underset{sp^3d^2}{XeF_5^{\oplus}} + \underline{F^-}$

$$I_{2(l)} \longrightarrow \underset{sp^3}{I_3^{\oplus}} + \underset{sp^3d}{I_3^{\ominus}}$$

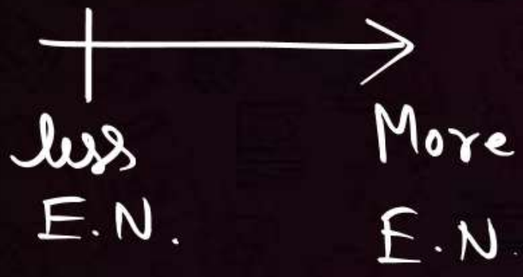
Dipole Moment

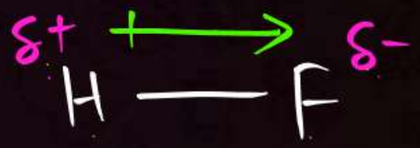
Peter Debye, the Dutch chemist received Nobel prize in 1936 for his work on X-ray diffraction and dipole moments. The magnitude of the dipole moment is given in Debye units in order to honour him.

$$\mu = q \times d$$

q = charge fraction

d = distance b/w charge / B.L.





Unit :→ Debye | C.m | esu-cm

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

$1D = 3.3 \times 10^{-30}$ C.m

NCERT CORNER [Page No:- 112]

Table 4.5 Dipole Moments of Selected Molecules

Type of Molecule	Example	Dipole Moment, μ(D)	Geometry
Molecule (AB)	HF	1.78	linear
	HCl	1.07	linear
	HBr	0.79	linear
	HI	0.38	linear
	H_2	0	linear
Molecule (AB_2)	H_2O	1.85	bent
	H_2S	0.95	bent
	CO_2	0	linear
Molecule (AB_3)	NH_3	1.47	trigonal-pyramidal
	NF_3	0.23	trigonal-pyramidal
	BF_3	0	trigonal-planar
Molecule (AB_4)	CH_4	0	tetrahedral
	$CHCl_3$	1.04	tetrahedral
	CCl_4	0	tetrahedral

Applications of Dipole Moment

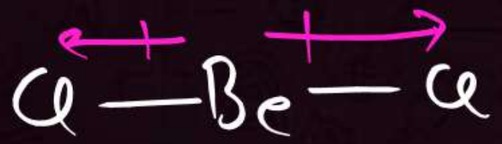
❖ **Polarity of Molecules** $\overset{\circ}{\underset{\circ}{\longrightarrow}}$

$\mu_r = 0 \Rightarrow$ N.P.

$\mu_r \neq 0 \Rightarrow$ Polar.

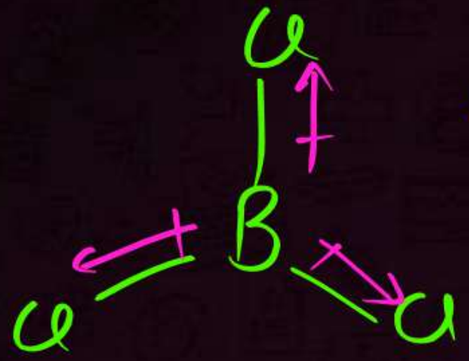
$\mu_{lp} \gg \mu_{bp}$

Ex: $BeCl_2$



$\mu_r = 0$ (N.P.)

BCl_3



$\mu_R = 0$

N.P.

* Hitrick * :→ 1 lp ⇒ Polar

Terminal atoms are same

- 0 lp ⇒ N.P.
- 1 lp ⇒ Polar
- 2 lp ⇒ sp^3d^2, sp^3d^3 ⇒ N.P.
  sp^3, sp^3d, ... ⇒ P.
- 3 lp ⇒ sp^3d ⇒ N.P.
  Otherwise ⇒ P.

N.P.	N.P.			
$BeCl_2$	BeF_2			
CCl_4	BCl_3	BF_3	BBr_3	
CF_4	CH_4	PCl_5	PF_5	SF_6
IF_7	SeF_6			

NF_3, NH_3, NCl_3, PH_3, PCl_3 SF_4 SeF_4 $SeCl_4$

XeO_2F_2 SO_2

1 lp $\Rightarrow$ Polar

$H_2\ddot{O}$: $\ddot{O}F_2$ $\ddot{Cl}F_3$

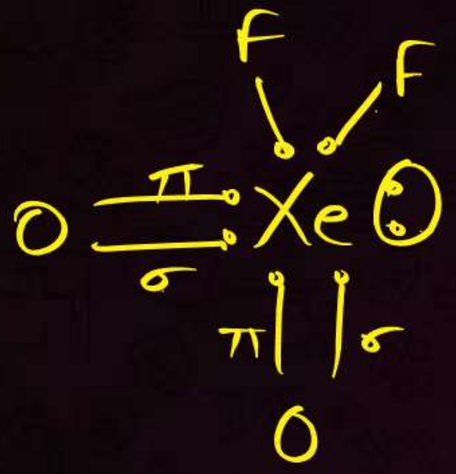
sp^3 sp^3 sp^3d

Polar

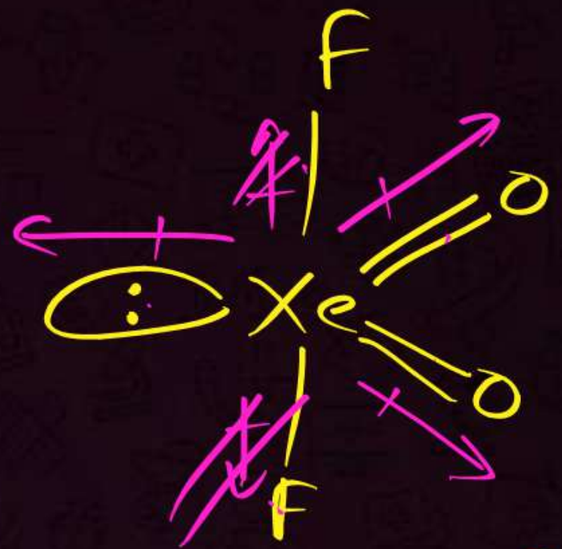
XeF_4 $\ddot{I}Cl_4^-$ $\ddot{I}F_4^-$

sp^3d^2

N.P.



sp^3d



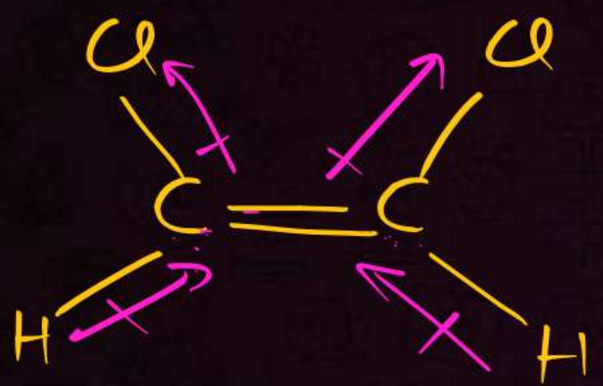
μ_{up}

μ_{down}

$\mu_R \neq 0$

Polar

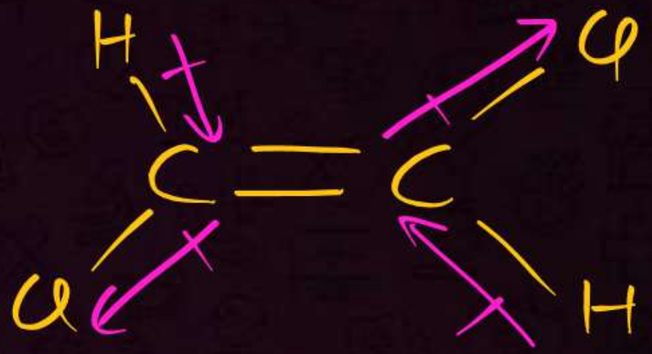
❖ **Distinction in cis and trans isomers**



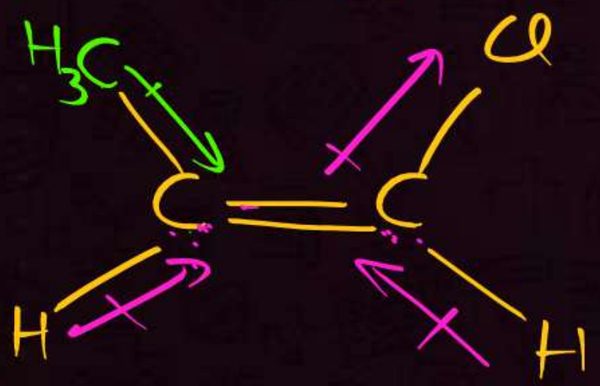
$\mu_R \neq 0$

Polar

>



$\mu_R = 0$ (N. P.)



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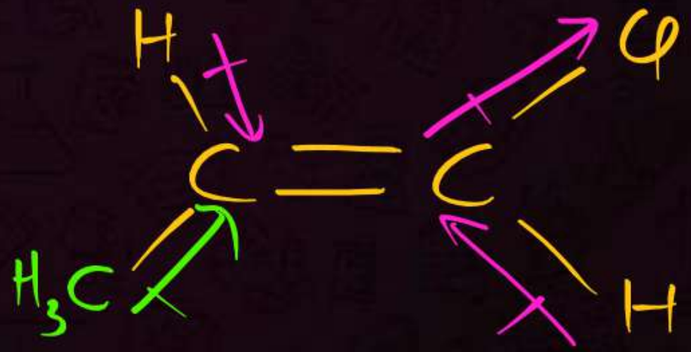
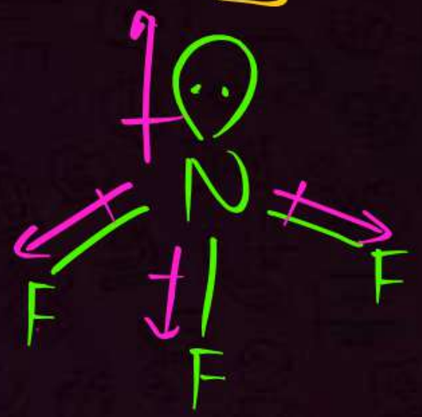


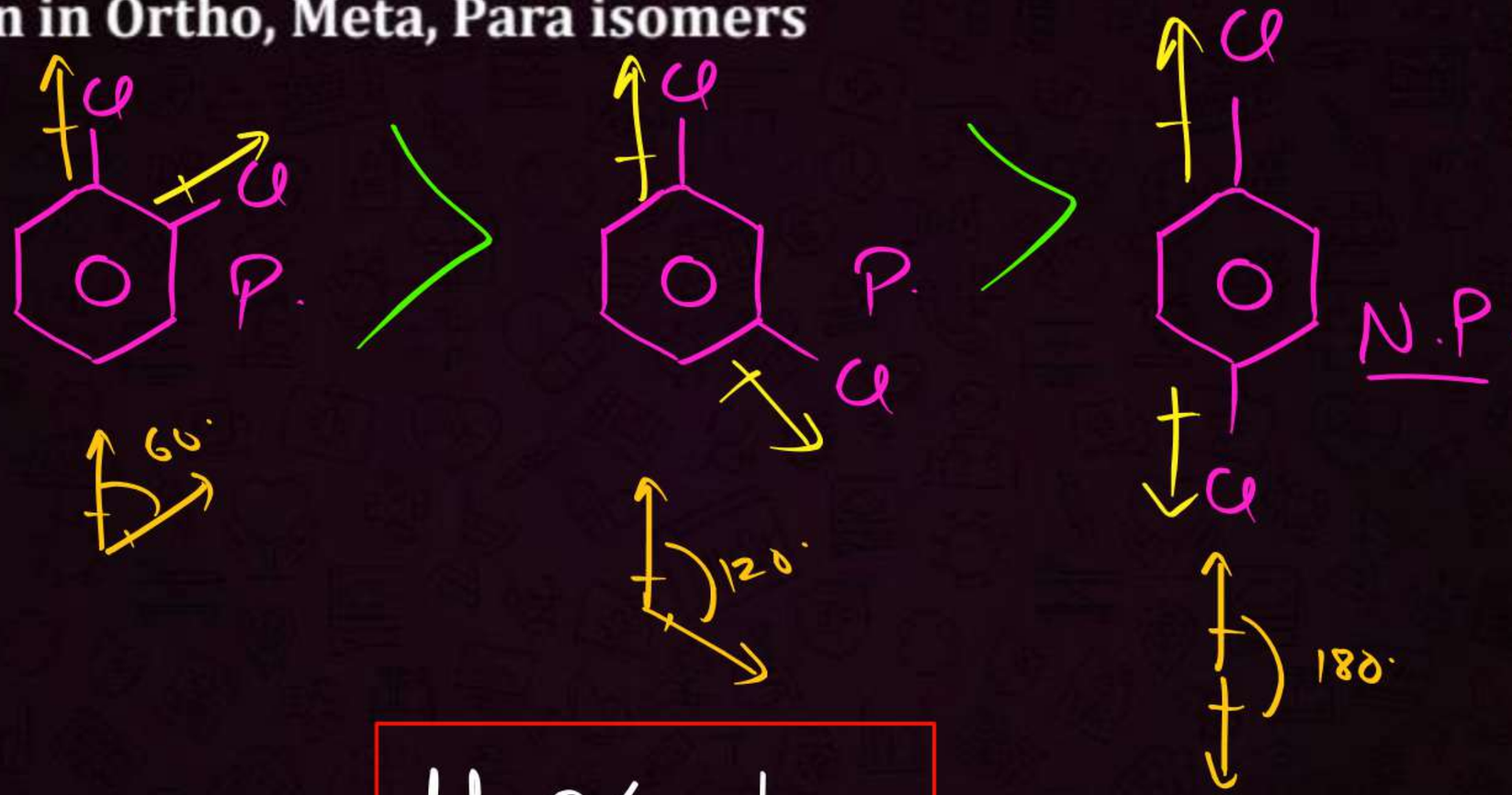
(add)

NH_3 > NF_3



More Polar

❖ **Distinction in Ortho, Meta, Para isomers**

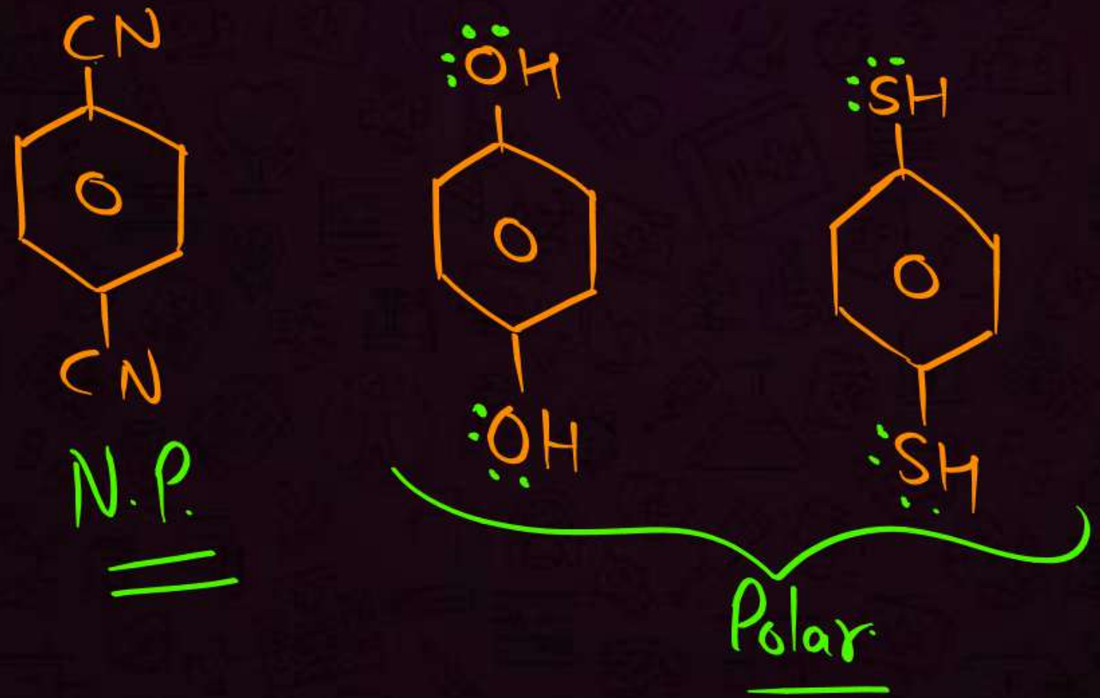


$$\mu_R \propto \frac{1}{B.A.}$$

* Remember *

① $CH_3Cl > CH_3F > CH_3Br > CH_3I$

② $\underbrace{CH_4 = CCl_4}_{N.P.} < CHCl_3 < CH_2Cl_2 < CH_3Cl$



QUESTION

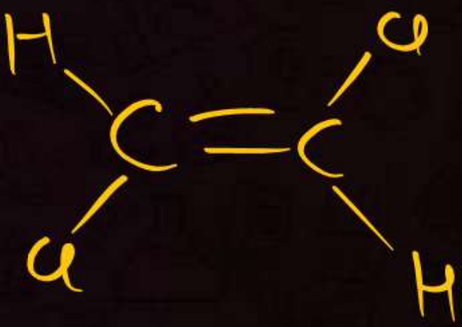
Of the following compounds, which will have a zero dipole moment? **(1987, 1M)**

(A) 1, 1-dichloroethylene

(B) cis-1, 2-dichloroethylene

(C) trans-1, 2-dichloroethylene

(D) None of the above



QUESTION

H.W.

Arrange the following compounds in order of increasing dipole moment, toluene (I), m-dichlorobenzene (II), 0-dichlorobenzene (III), p-dichlorobenzene (IV)

(1996. 1M)

(A) I < IV < II < III

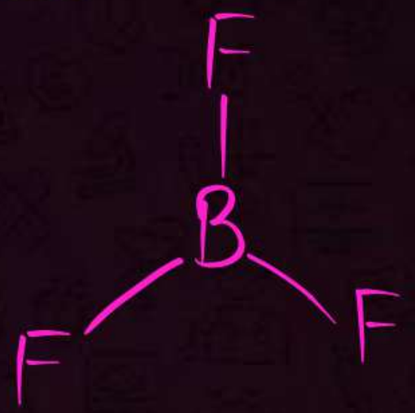
(B) IV < I < II < III

(C) IV < I < III < II

(D) IV < II < I < III

QUESTION

The molecule which has zero dipole moment is (1989, 1M)



(A) CH_2Cl_2

(B) BF_3

(C) NF_3

(D) ClO_2

QUESTION

Among the following, the molecule with the highest dipole moment is

(2003, 1M)

A CH_3Cl

B CH_2Cl_2

C $CHCl_3$

D CCl_4

❖ % Ionic Character

$$\% \text{ I.C.} = \frac{\mu_{obs.}}{\mu_{calc.}} \times 100$$

Ionic Bond

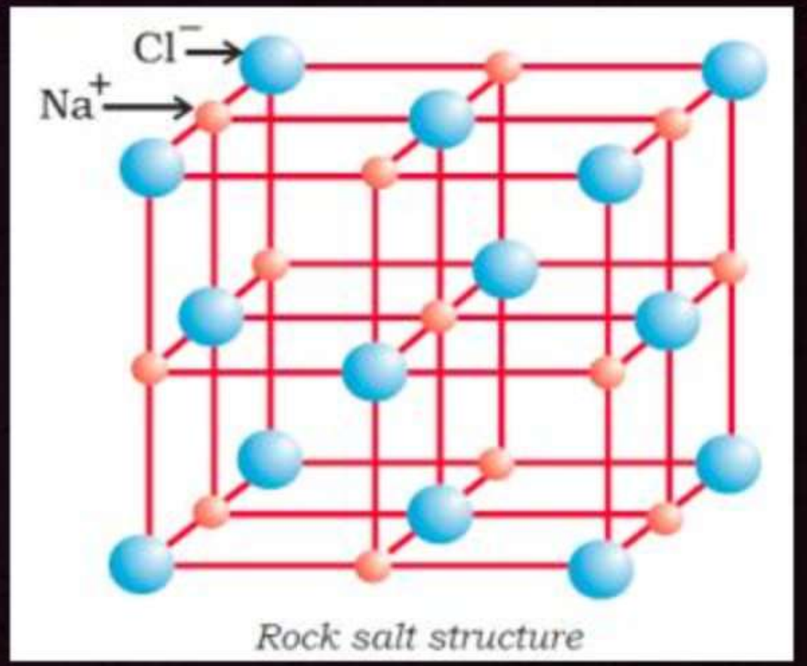
"Electrovalent Bond"

↳ electrostatic force of attⁿ.

NCERT CORNER [Page No:- 106]

From the Kössel and Lewis treatment of the formation of an ionic bond, it follows that the formation of ionic compounds would primarily depend upon:

- The ease of formation of the positive and negative ions from the respective neutral atoms;
- The arrangement of the positive and negative ions in the solid, that is, the lattice of the crystalline compound.



Rock salt structure

* Lattice Energy :↓

The Lattice Enthalpy of an ionic solid is defined as the energy required to completely separate one mole of a solid ionic compound into gaseous constituent ions.

$$NaCl_{(s)} \xrightarrow{L.E.\downarrow} Na^{+}_{(g)} + Cl^{-}_{(g)}$$

Note: **[Page No:- 107]**

Thus a qualitative measure of the stability of an ionic compound is provided by its enthalpy of lattice formation and not simply by achieving octet of electrons around the ionic species in gaseous state.

QUESTION

The number and type of bonds between two carbon atoms in CaC_2 are (1996, 1M)

A one sigma (σ) and one pi (π) bonds

B one sigma (σ) and two pi (π) bonds

C one sigma (σ) and one half pi (π) bonds

D one sigma (σ) bond

Lattice Enthalpy

The amt. of energy released during formation of one mole ionic crystal lattice from its constituent ions.

$$Na^+_{(g)} + Cl^-_{(g)} \xrightarrow{E\uparrow} NaCl_{(s)}$$

or

The amt. of energy required to break one mole ionic crystal lattice into its constituent ions.

$$NaCl_{(s)} \xrightarrow{E\downarrow} Na^+_{(g)} + Cl^-_{(g)}$$

Factors Affecting L.E.

① Charge :→ $L.E. \propto |q_1 q_2|$

② Size :→ $L.E. \propto \frac{1}{\text{size}}$ $\quad \frac{1}{(r_+ + r_-)}$

Ex: $\overset{+1}{Na}\overset{-1}{F} < \overset{+2}{Mg}\overset{-1}{F_2} < \overset{+3}{Al}\overset{-1}{F_3}$

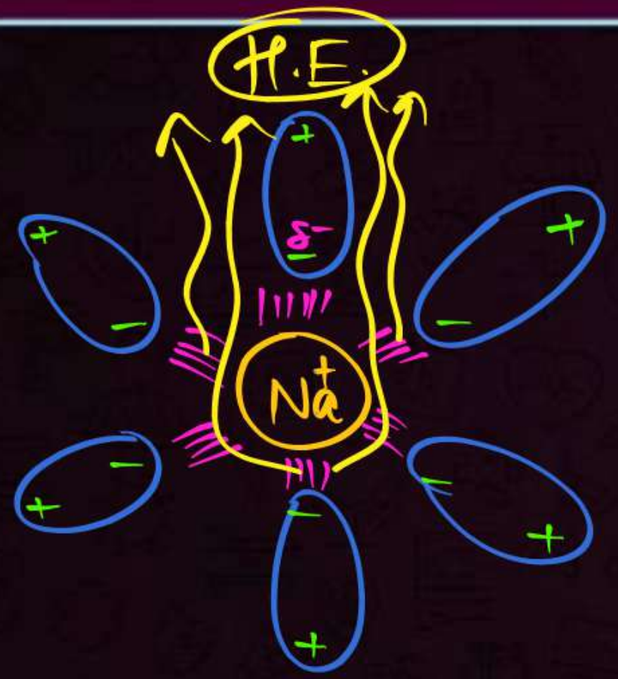
$\overset{+1}{Li}\overset{-1}{F} > \overset{+1}{Na}\overset{-1}{F} > \overset{+1}{K}\overset{-1}{F} > \overset{+1}{Rb}\overset{-1}{F} > \overset{+1}{Cs}\overset{-1}{F}$

$\overset{+3\ -1}{AlF_3} < \overset{+3\ -2}{Al_2O_3} < \overset{+3\ -3}{AlN}$

$\overset{+2\ -1}{MgF_2} < \overset{+2\ -2}{MgO} < \overset{+2\ -3}{Mg_3N_2}$

$\overset{+2\ -2}{BeO} > \overset{+2\ -2}{MgO} > \overset{+2\ -2}{CaO} > \overset{+2\ -2}{SrO} > \overset{+2\ -2}{BaO}$

Hydration Enthalpy



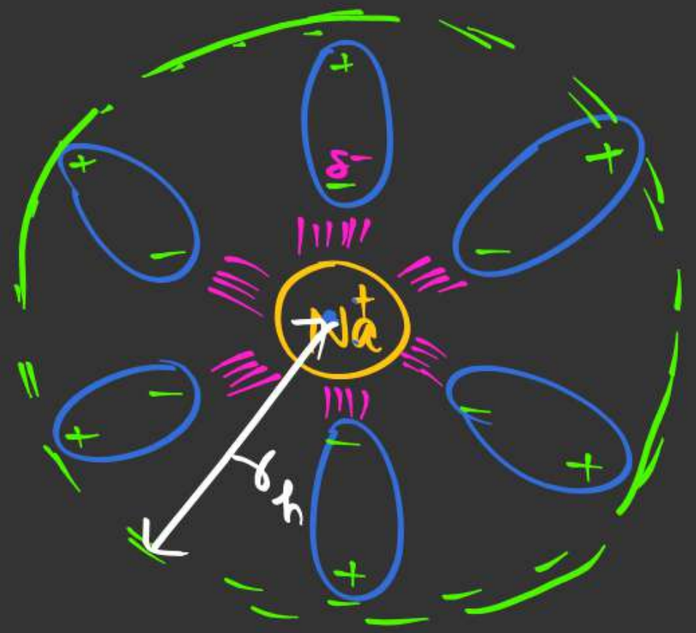
① Charge ⇒ H.E ∝ charge

② Size ⇒ $H.E \propto \frac{1}{size}$

Ex: $Na^+ < Mg^{+2} < Al^{+3}$

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

Hydrated Rad. $\propto$ H.E.
(r_h)



Size: ↴

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

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

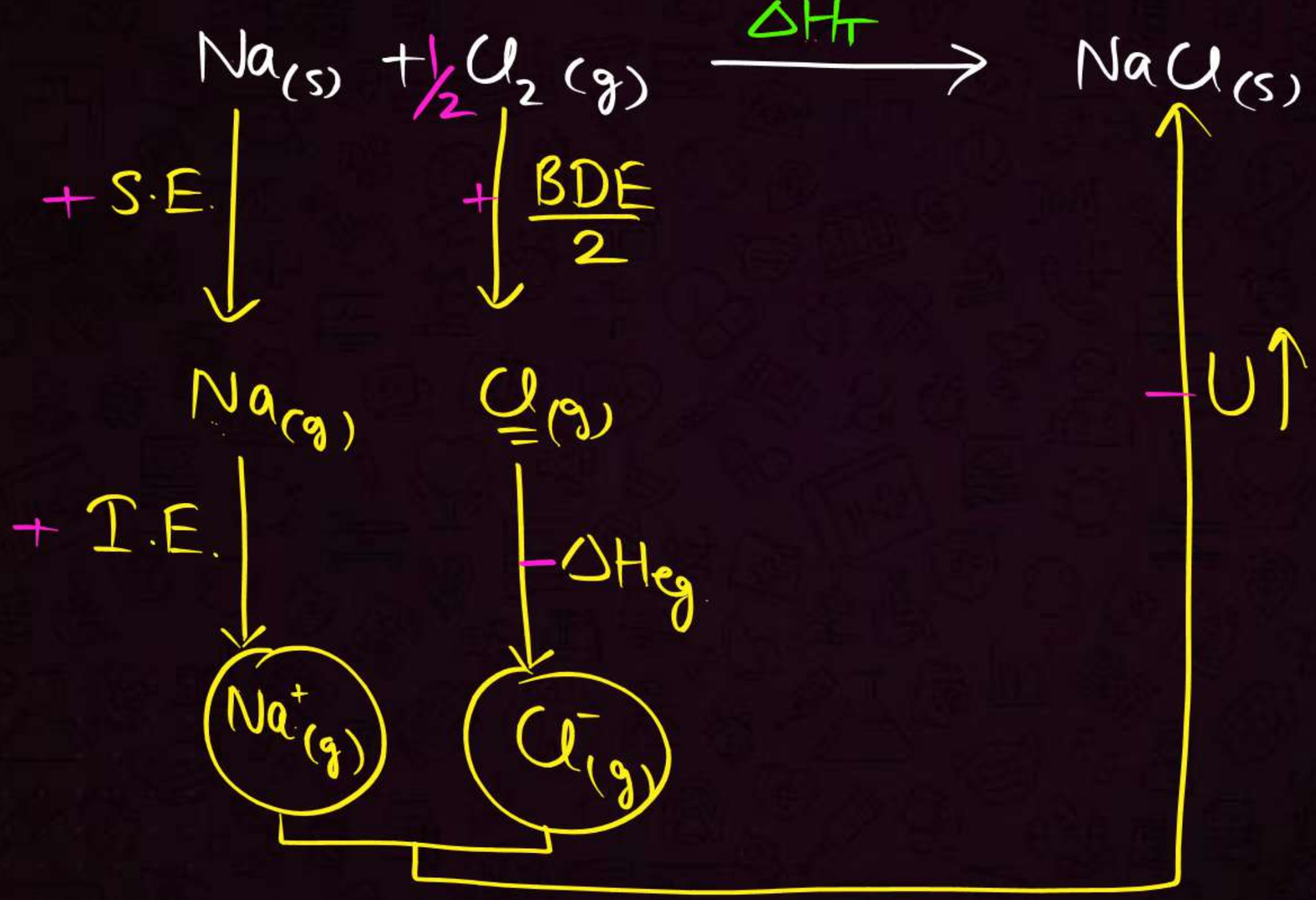
Mobility $\propto \frac{1}{r_h}$

Mobility : $Li^+_{(aq.)} < Na^+_{(aq.)} < K^+_{(aq.)} < Rb^+_{(aq.)} < Cs^+_{(aq.)}$

conductivity $\propto$ Mobility

conductivity $\therefore \rightarrow$ $Li^+_{(aq.)} < Na^+_{(aq.)} < K^+_{(aq.)} < Rb^+_{(aq.)} < Cs^+_{(aq.)}$

Born Haber's Cycle

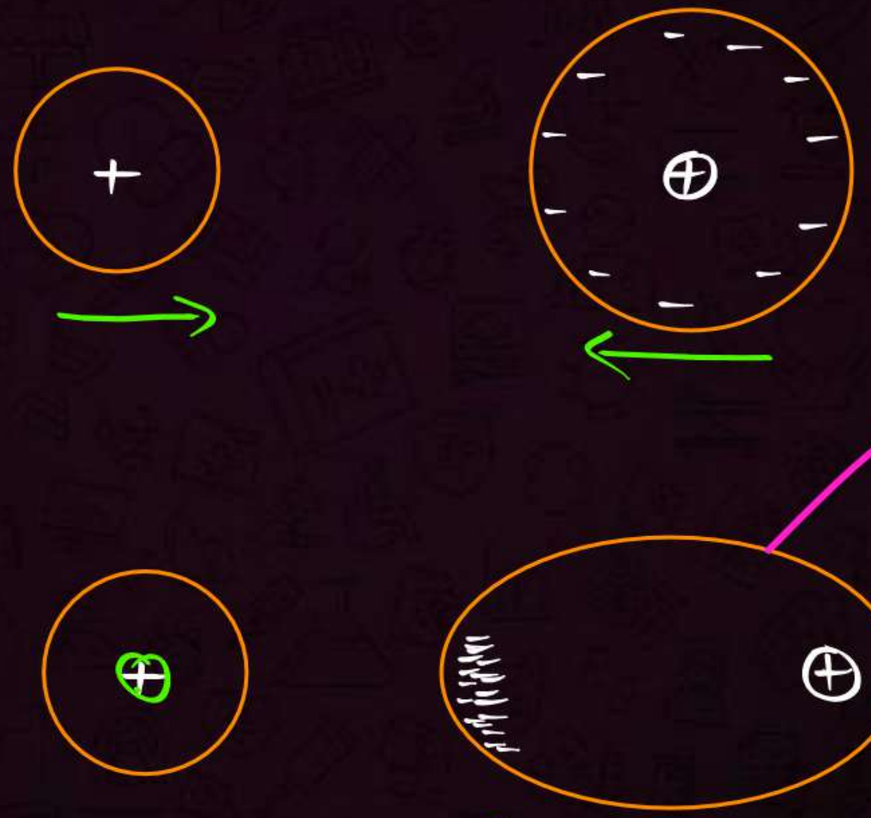


$$\Delta H_T = S.E + I.E + \frac{BDE}{2} - \Delta H_{eg} - U$$

formation of NaCl $\Rightarrow \Delta H_T = -ve$

$$SE + IE + \frac{BDE}{2} < -\Delta H_{eg} - U$$

Fajan's Rule (Polarisation)



Cov. Nature & Polarization

Polarization depends on :

1. Polarising Power of Cation (ϕ)
2. Polarizability of Anion

for Anion (Polarizability)

① Charge of Anion ⇒ Pol α charge

② Size of Anion ⇒ Pol α size

Ex: $F^- < Cl^- < Br^- < I^-$

$F^- < O^{-2} < N^{-3} < C^{-4}$

for Cations (Pol. Power)

① charge ⇒ $\phi \propto$ charge

② size ⇒ $\phi \propto \frac{1}{\text{size}}$

Ex: $Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$

$Na^+ < Mg^{+2} < Al^{+3} < Si^{+4}$

Covalent Nature :↓

$$LiF > NaF > KF > RbF > CsF$$

$$LiF < Li_2O < Li_3N$$

$$LiF < LiCl < LiBr < LiI$$

$$SnCl_2 < SnCl_4$$

$$PbCl_2 < PbCl_4$$

③ e⁻ config of Cation :-

$$8e^- < 18e^- < (18+2)e^-$$

Na^+, Ca^{+2}
Mg^{+2}, K^+

Zn^{+2}
Cu^+
Ag^+
Cd^{+2}
Hg^{+2}
Ga^{+3}

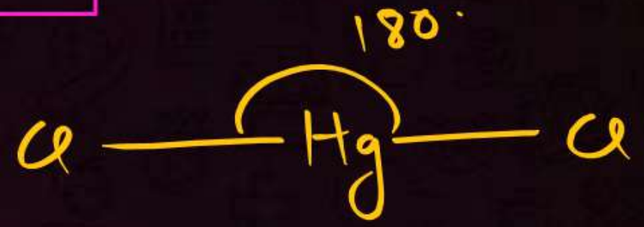
(Tl^+, Pb^{+2}, Bi^{+3})

$\overset{+2}{Ca}Cl_2 < \overset{+2}{Zn}Cl_2$

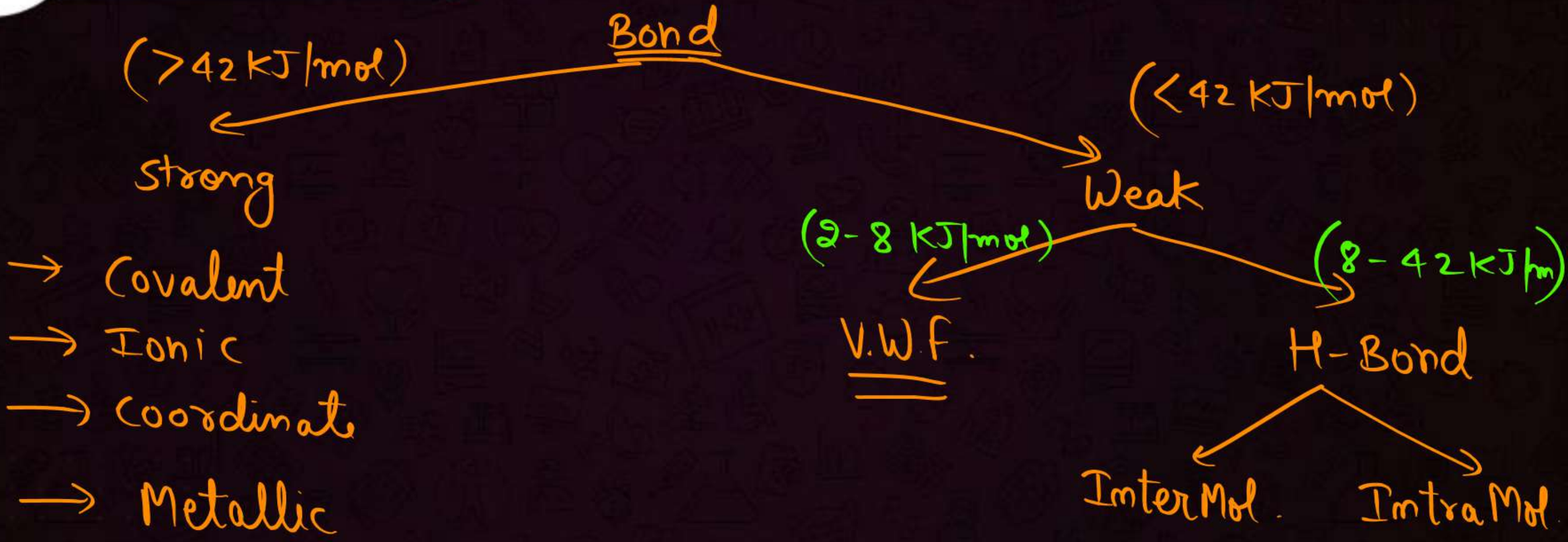
$\overset{+}{K}Cl < CuCl$

$$Zn^{+2}Cl_2 < Cd^{+2}Cl_2 < Hg^{+2}Cl_2$$

Cov. Nature $\longrightarrow$



Weak forces


Bond
(>42 KJ/mol)
(<42 KJ/mol)
Strong
Weak
→ Covalent
→ Ionic
→ Coordinate
→ Metallic
(2-8 KJ/mol)
(8-42 KJ/m
V.W.F.
H-Bond
InterMol.
IntraMol.

(Polar)

① Dipole-Dipole force (Keesom force) ⇒ (Polar + Polar)

Ex: $HCl + HCl$

$HCl + H_2O$

② Dipole-Induced Dipole (Debye forces)

(Polar + Non Polar)

$Cl_2 + HCl$

③ Instantaneous Dipole Induced Dipole (L.D.F.)

(Non Polar + Non Polar)

Ex: $Cl_2 + Cl_2$

① ion - Dipole ⇒ ion + Polar

$F^{\ominus} + HF$

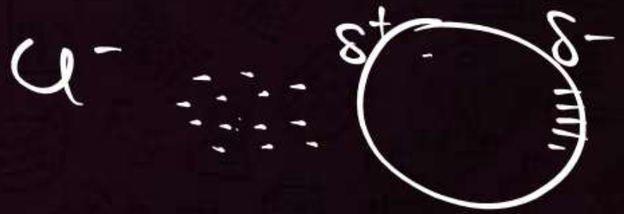
$HF_2^{\ominus}$

KHF_2

② Ion Induced Dipole

(ion + Non Polar)

$Cl^- + CCl_4$



P P
$HF + HF \Rightarrow$ D-D

$HCl + H_2O \Rightarrow$ DD.

$HF + HCl \Rightarrow$ D. D

$HCl + Cl_2$ DID.

$H_2O + CCl_4$ D.I.D.

$H_2O + C_6H_6$ DID

$C_6H_6 + C_6H_6$ IDID

$CCl_4 + C_6H_6$ IDID

$C_6H_6 + F^{\ominus}$ Ion-induced dipole

$C_6H_6 + CS_2$ (N.P.) IDID

$CS_2 + H_2O$ DID

$Xe + H_2O$ DID

$Xe + Xe$ IDID

$F^{\ominus}$ + HF — ion - dipole

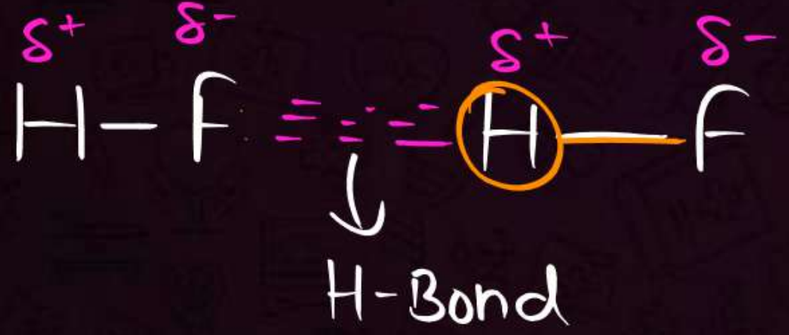
$Cl^{\ominus}$ + HCl — —"— —

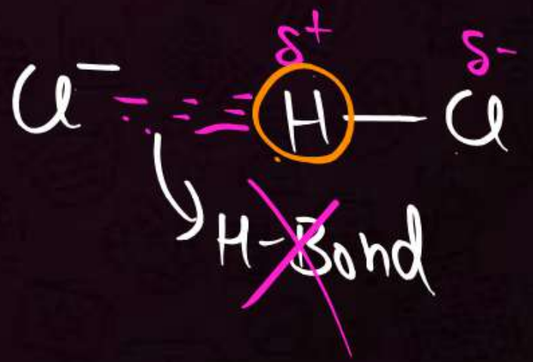
I^- + I_2 } ion - induced dipole

I^+ + I_2

Hydrogen Bonding

Type of ion-dipole &/or dipole dipole Interaction





Condition :

QUESTION

The maximum possible number of hydrogen bonds a water molecule can form is **(1992, 1M)**

A 2

B 4

C 3

D 1

QUESTION

Which one among the following does not have the hydrogen bond? **(1983, 1M)**

(A) Phenol

(B) Liquid NH_3

(C) Water

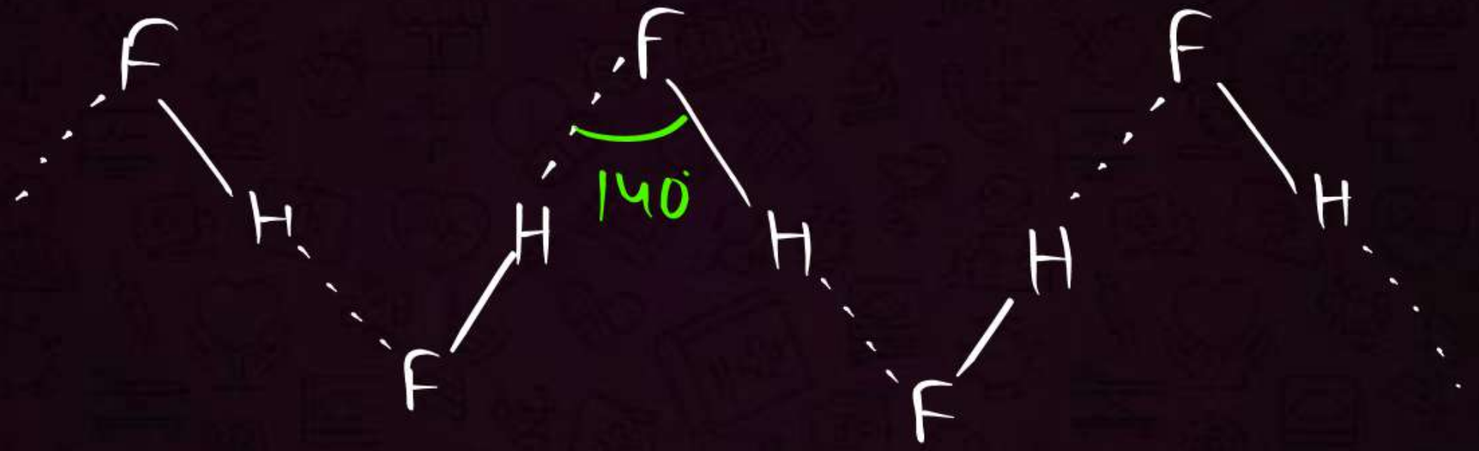
(D) HCl

Types of Hydrogen Bonding

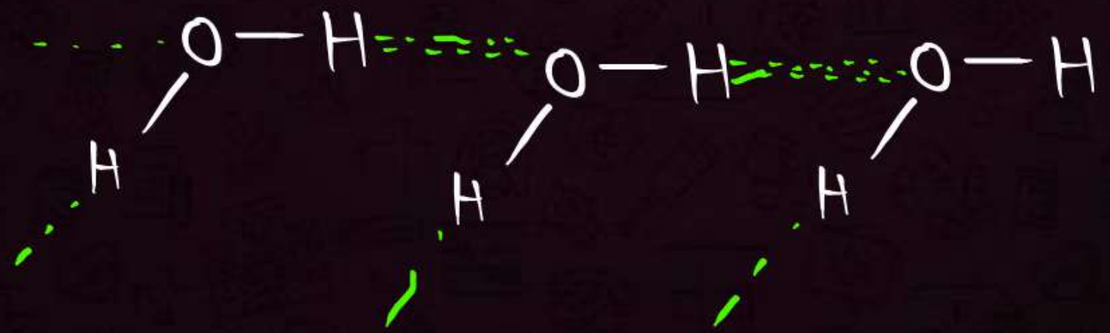
Intermolecular H-Bonding

(b/w the molecules)

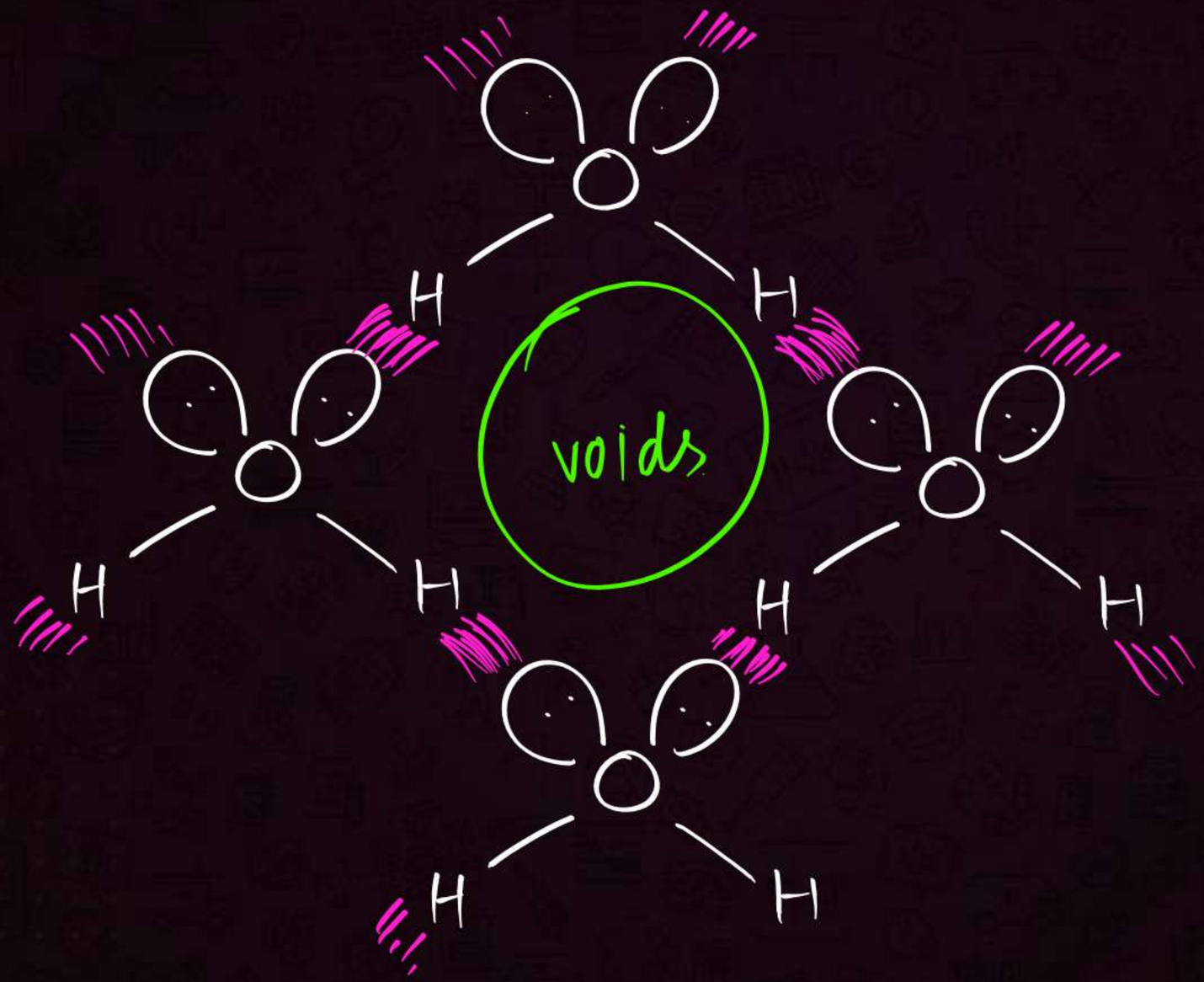
HF



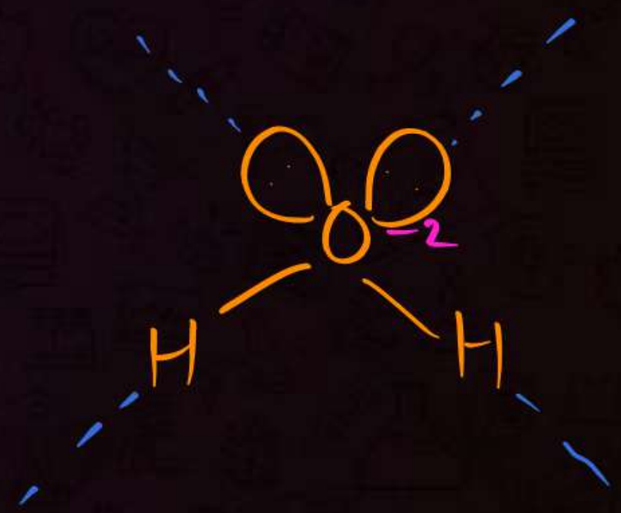
H_2O



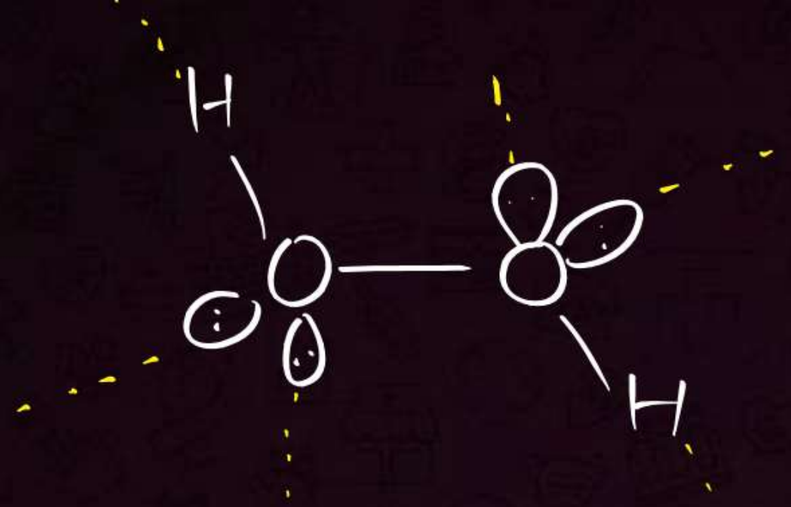
$H_2O_{(s)}$ (ice)



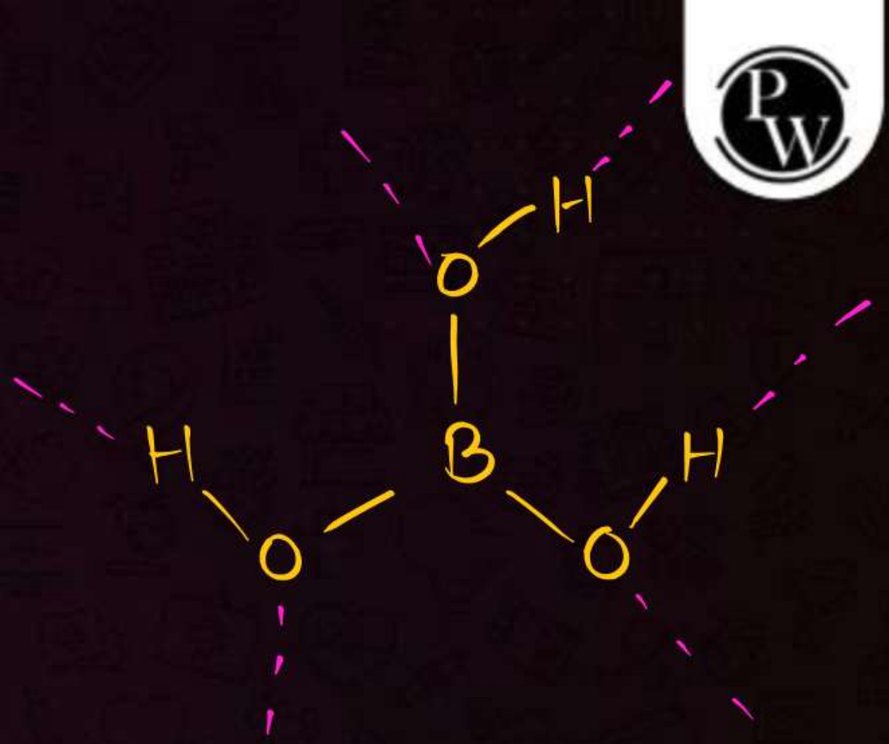
V↑, D↓



H_2O_2



no. of H-Bond = 6

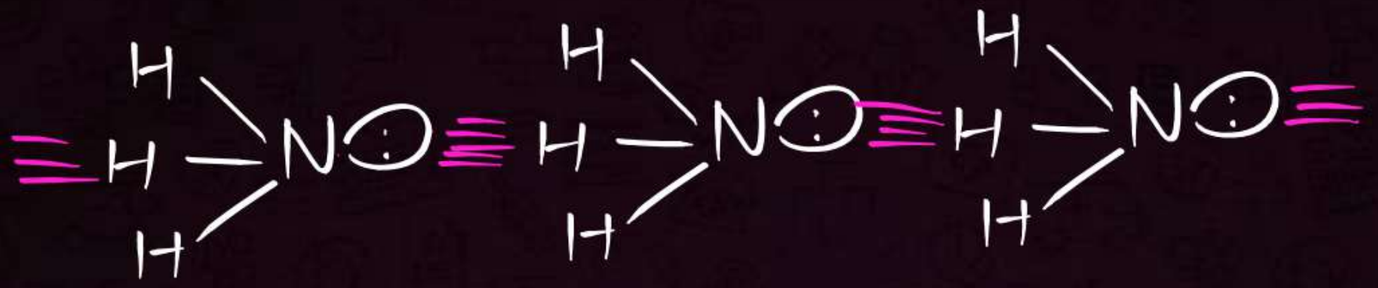


No. of H-Bonds = 6

Extent of H-Bond = $H_2O < H_2O_2$

Strength of H-Bond = $H_2O > H_2O_2$

NH_3

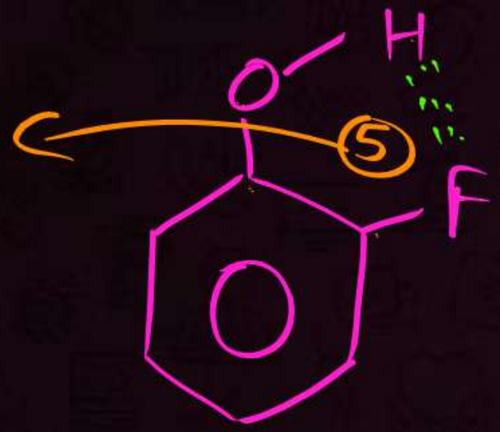


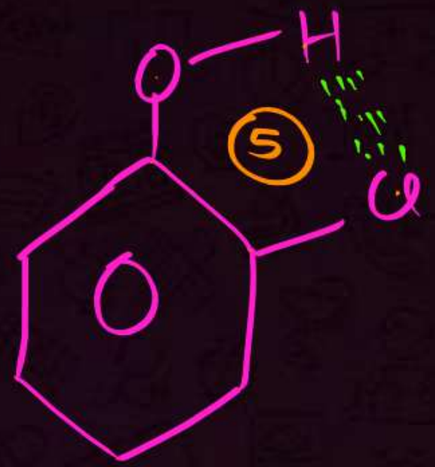
HCN

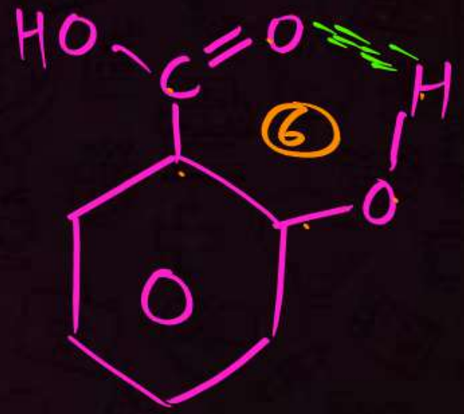
II) Intramolecular H-Bond :→

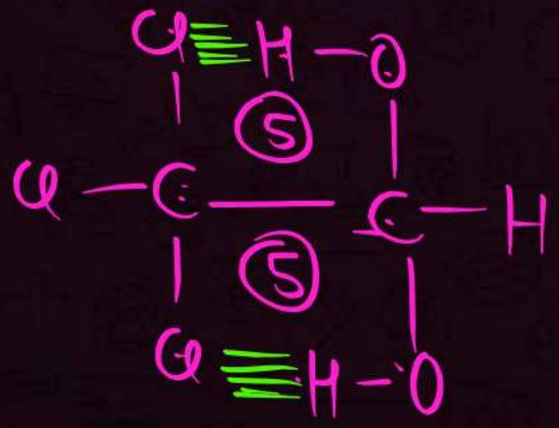
(within a Molecule)

Chelate Ring (5-6)

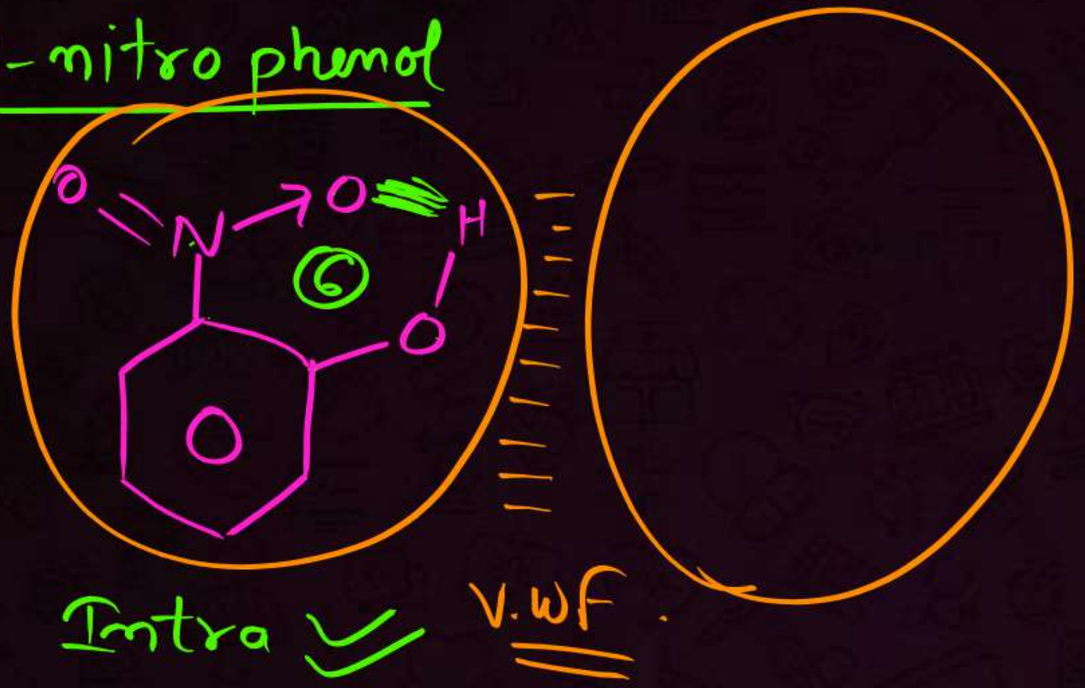








O-nitro phenol

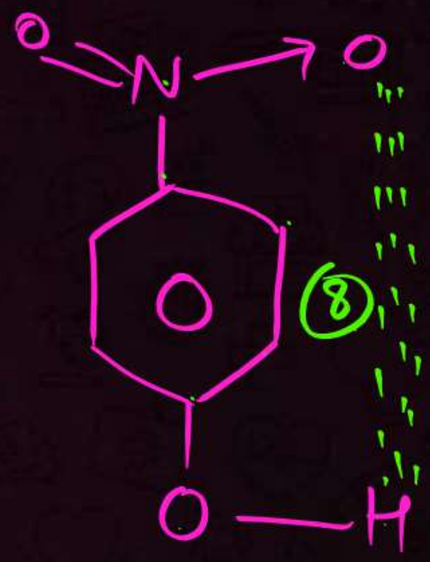


Intra > V.wf

P/Q

V.P. ⇒ O > P

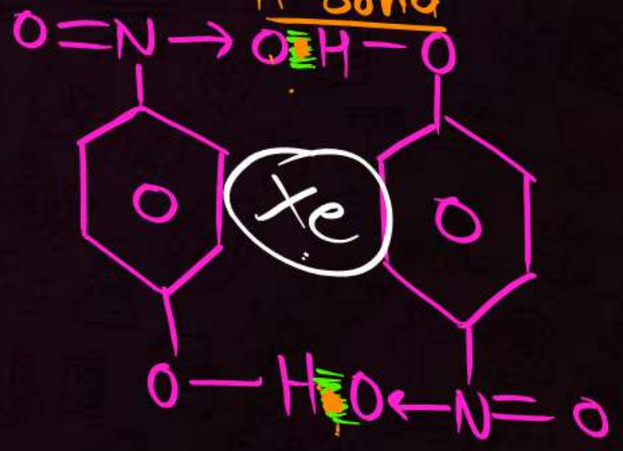
P-nitro phenol



X

Intra X

H-Bond



Clatherati

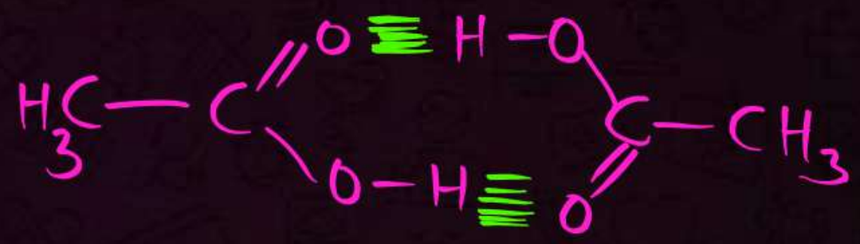
Application of Hydrogen Bonding

Properties :	Inter Mol.	Intra Mol
① B.P./M.P.	↑	↓
② Solubility	↑	↓
③ Volatility	↓	↑
④ Vapour pressure	↓	↑
⑤ Viscocity	↑	↓
⑥ Surface Tension	↑	↓

① Cellulose ⇒ intra Mol. H-Bonding

② DNA ⇒ H-Bonding

③ CH_3COOH



Molecular Orbital Theory

$e^- \rightarrow$ wave

(i) The electrons in a molecule are present in the various molecular orbitals as the electrons of atoms are present in the various atomic orbitals.

(ii) The atomic orbitals of comparable energies and proper symmetry combine to form molecular orbitals.

NEET

(iii) While an electron in an atomic orbital is influenced by one nucleus, in a molecular orbital it is influenced by two or more nuclei depending upon the number of atoms in the molecule. Thus, an atomic orbital is monocentric while a molecular orbital is polycentric.

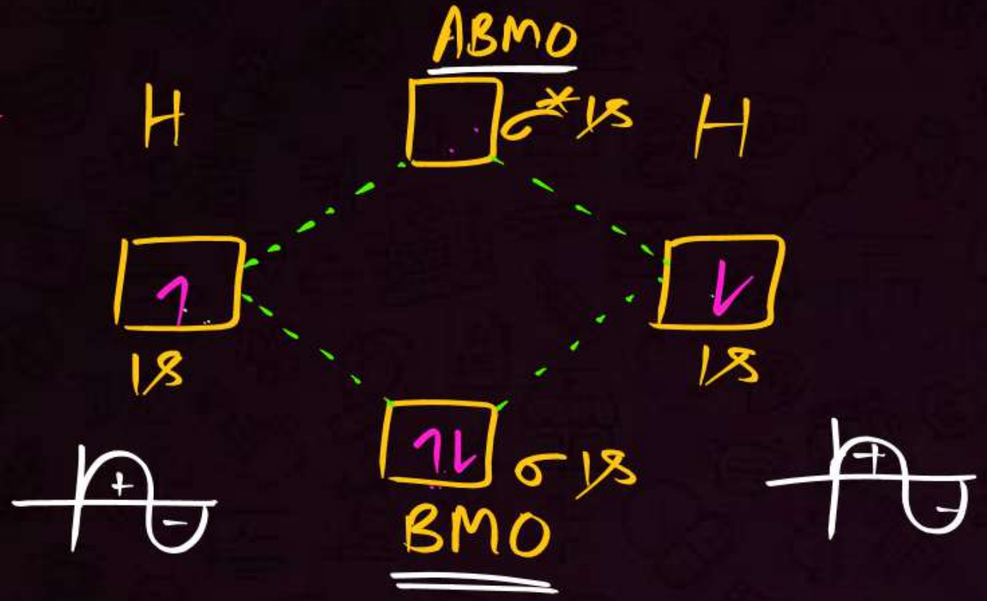
(iv) The number of molecular orbital formed is equal to the number of combining atomic orbitals. When two atomic orbitals combine, two molecular orbitals are formed. One is known as bonding molecular orbital while the other is called antibonding molecular orbital.

(v) The bonding molecular orbital has lower energy and hence greater stability than the corresponding antibonding molecular orbital.

(vi) Just as the electron probability distribution around a nucleus in an atom is given by an atomic orbital, the electron probability distribution around a group of nuclei in a molecule is given by a molecular orbital.

(vii) The molecular orbitals like atomic orbitals are filled in accordance with the aufbau principle obeying the Pauli's exclusion principle and the Hund's rule.

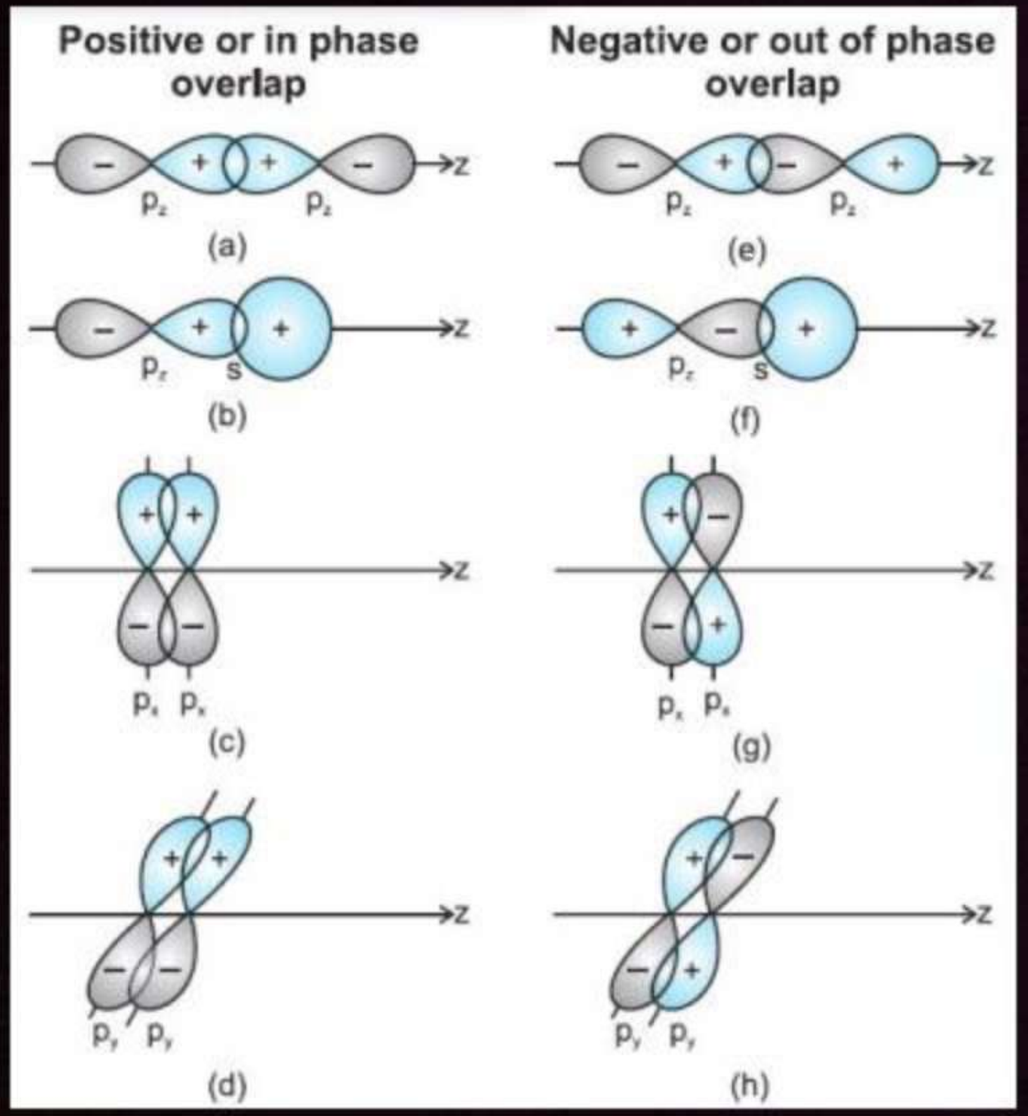
$H_2 = 2e^-$

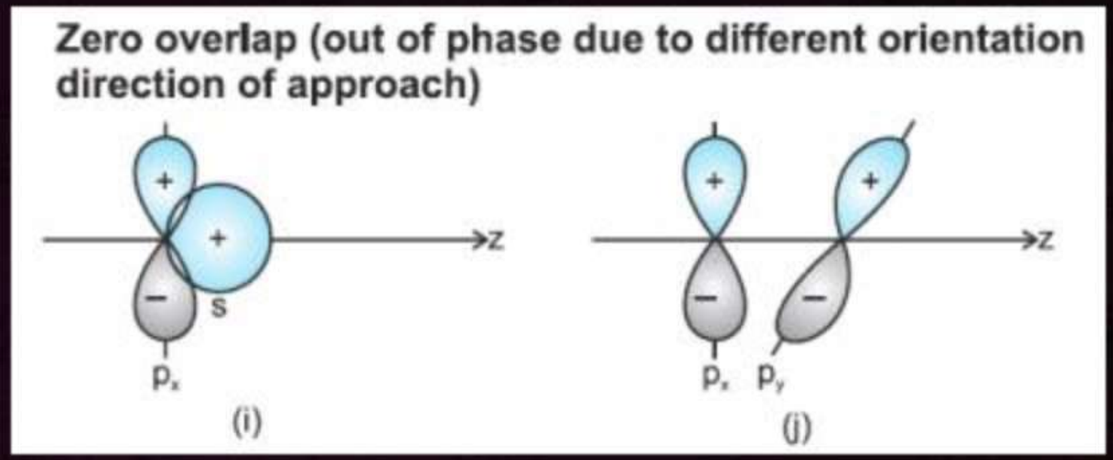


$$B.O. = \frac{1}{2}(B_{e^-} - AB_{e^-})$$

$$B.O. = \frac{2-0}{2} = 1$$

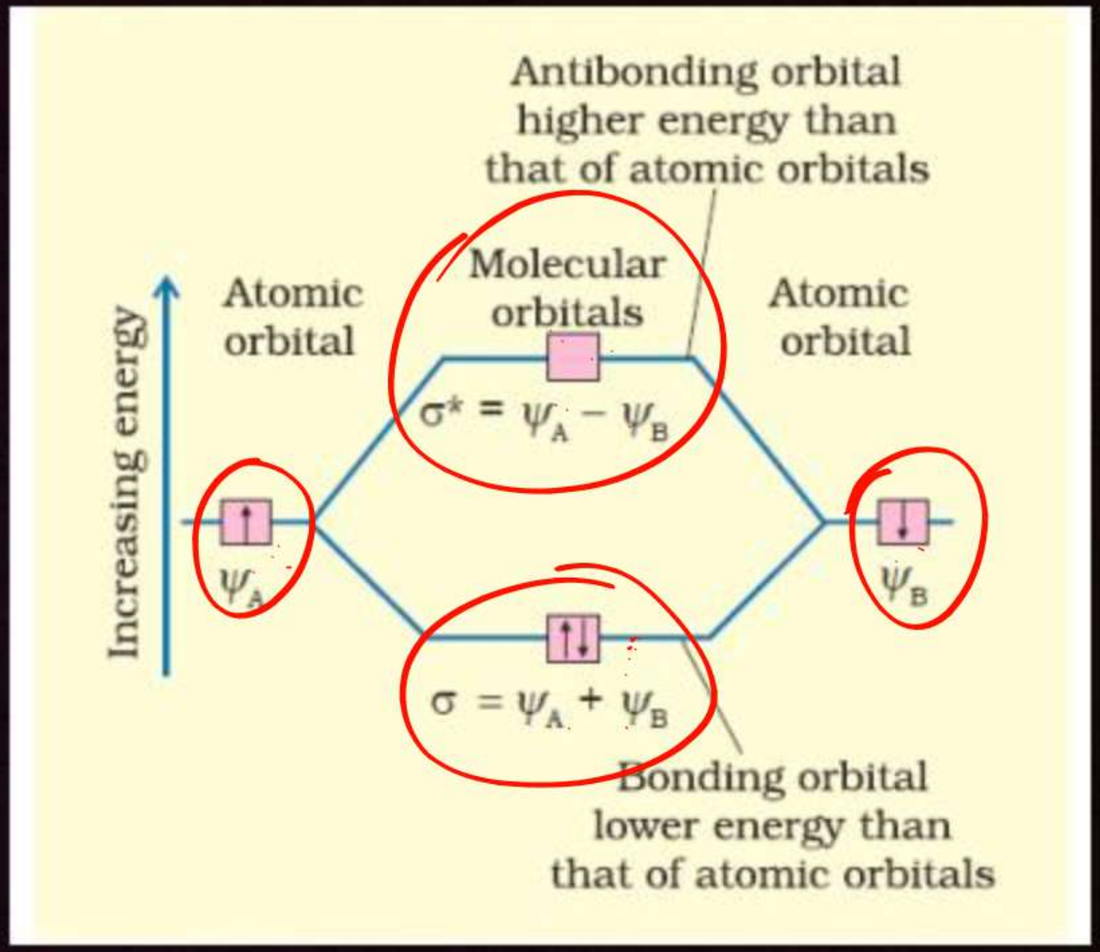
NCERT CORNER [Page No:- 119]





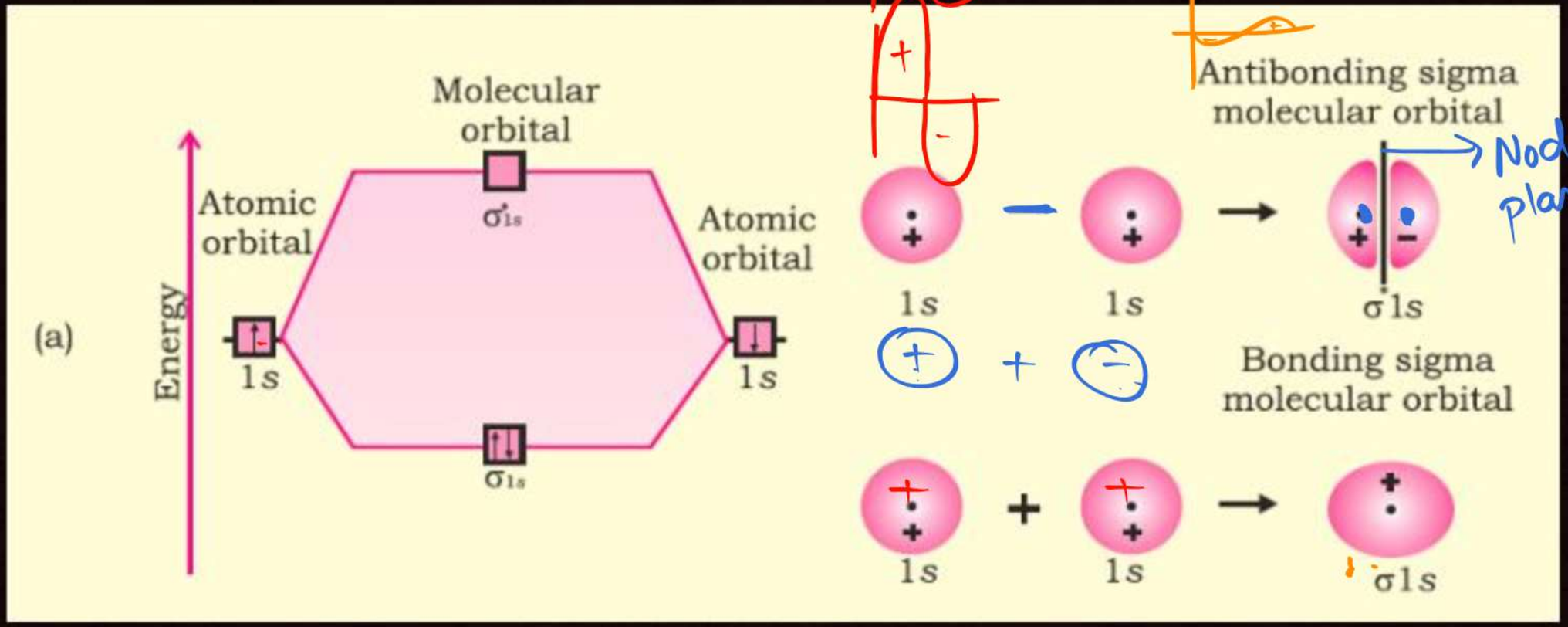
NCERT CORNER [Page No:- 126]

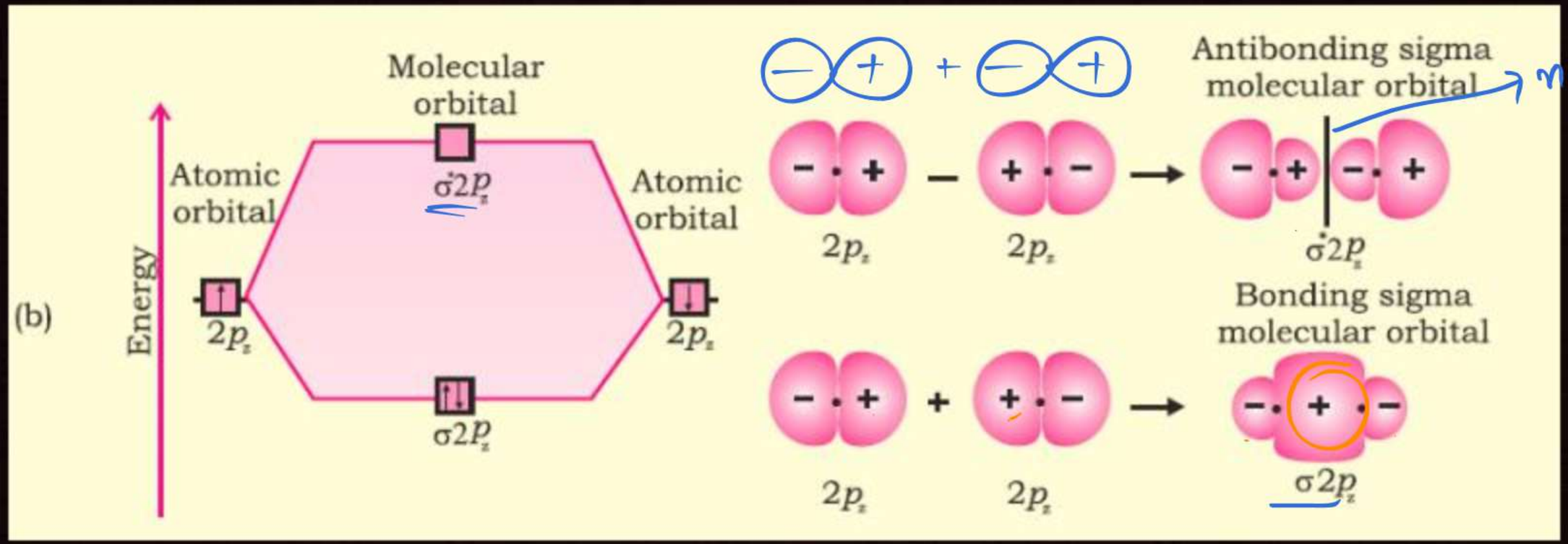
The molecular orbital σ formed by the addition of atomic orbitals is called the **bonding molecular orbital** while the molecular orbital σ* formed by the subtraction of atomic orbital is called **antibonding molecular orbital** as depicted in Fig. 4.19.

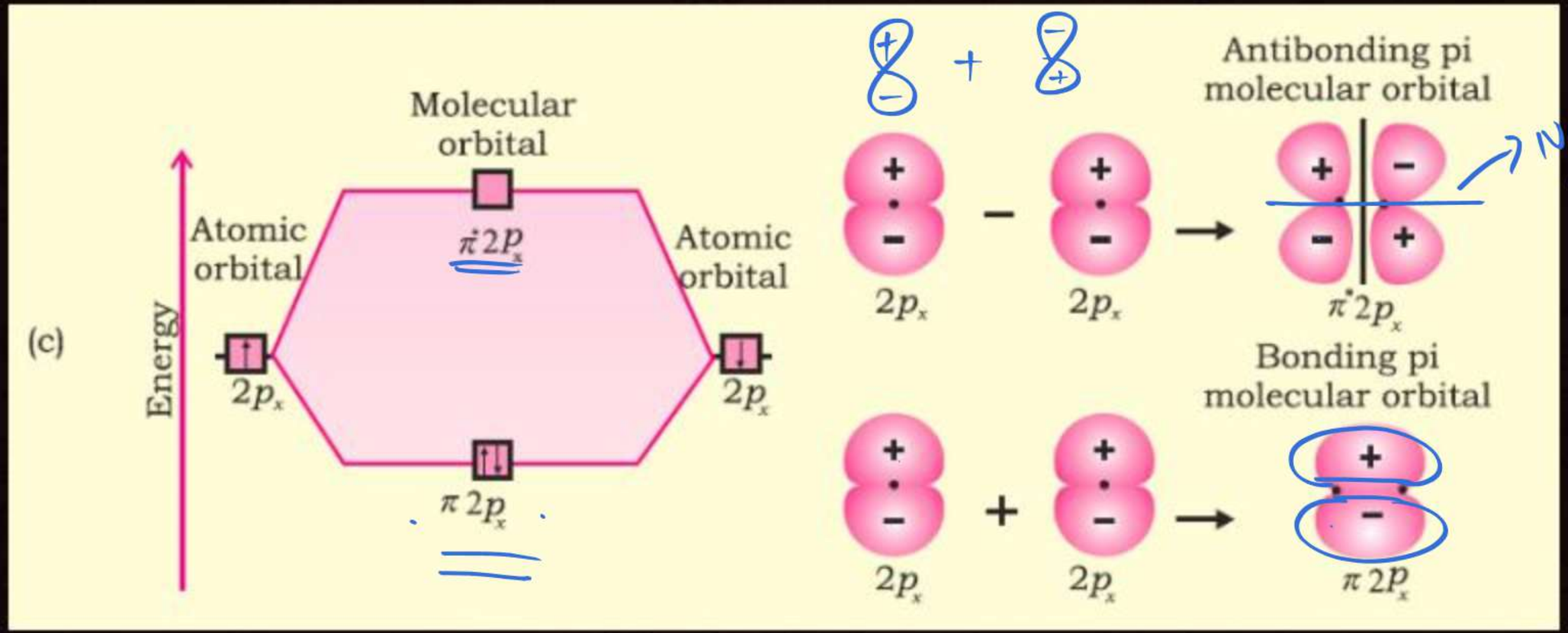


BMO < AO < ABMO

Energy




(b)
Energy
Molecular orbital
Atomic orbital
Atomic orbital
σ2p_z
2p_z
2p_z
σ2p_z
Antibonding sigma molecular orbital
nodal plane
2p_z
2p_z
σ2p_z
Bonding sigma molecular orbital
2p_z
2p_z
σ2p_z


(c)
Energy
Molecular orbital
Atomic orbital
Atomic orbital
$\pi^* 2p_x$
$2p_x$
$2p_x$
$\pi 2p_x$
Antibonding pi molecular orbital
$2p_x$
$2p_x$
$\pi^* 2p_x$
Nodal plane
Bonding pi molecular orbital
$2p_x$
$2p_x$
$\pi 2p_x$

NCERT CORNER [Page No:- 129]

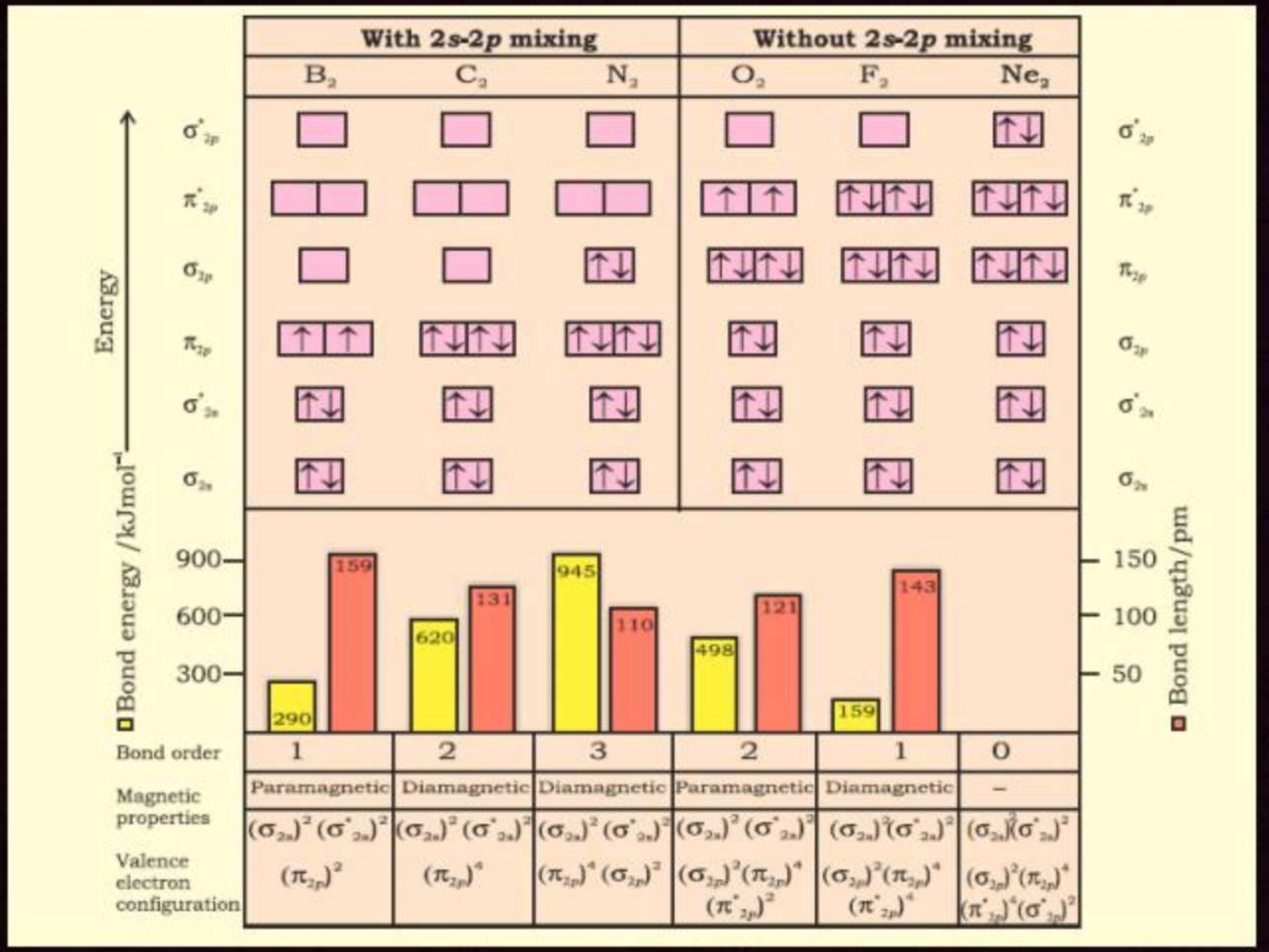
4.7.5 Electronic Configuration and Molecular Behaviour

The distribution of electrons among various molecular orbitals is called the **electronic configuration of the molecule**. From the electronic configuration of the molecule, it is possible to get important information about the molecule as discussed below.

Stability of Molecules: If N_b is the number of electrons occupying bonding orbitals and N_a the number occupying the antibonding orbitals, then

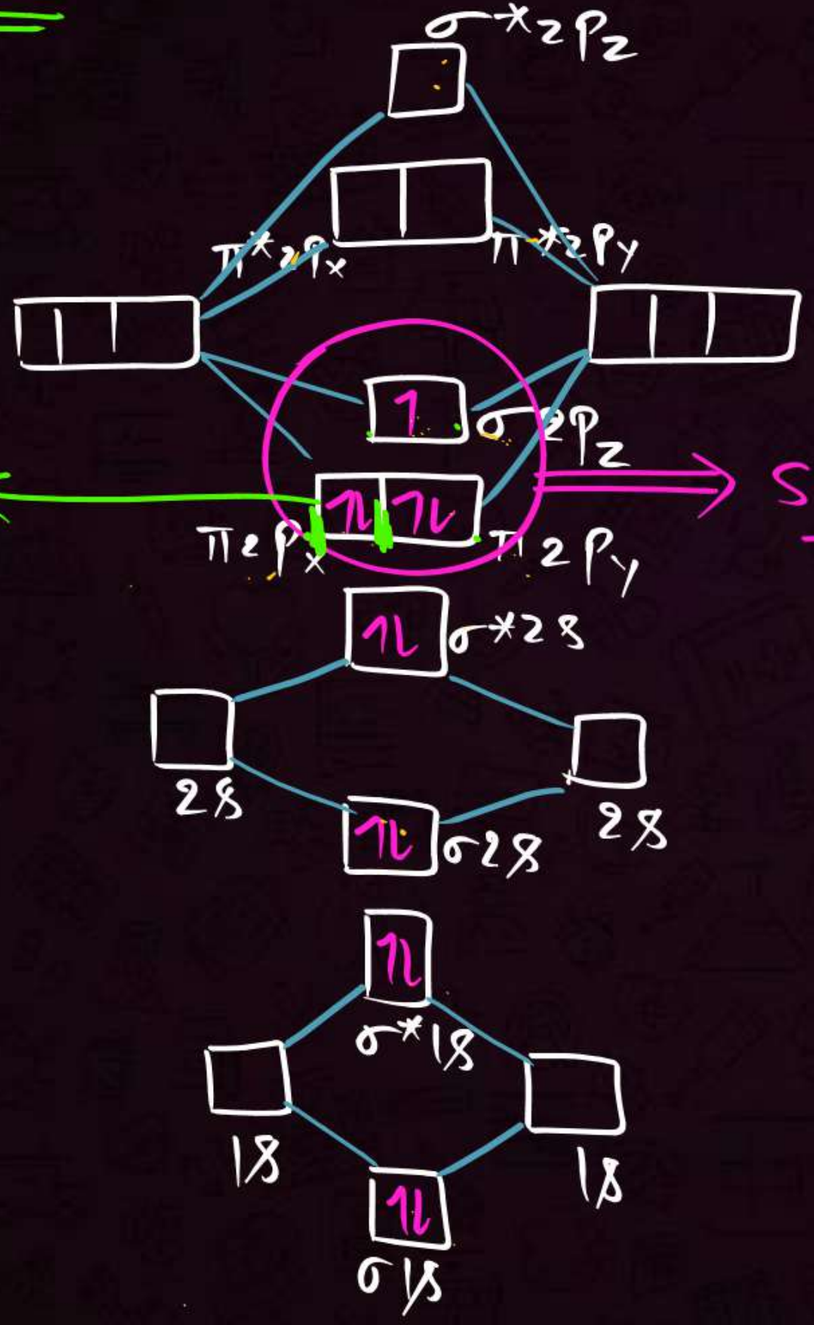
(i) the molecule is stable if N_b is greater than N_a, and

(ii) the molecule is unstable if N_b is less than N_a.


With 2s-2p mixing
Without 2s-2p mixing
B2
C2
N2
O2
F2
Ne2
σ*2p
π*2p
σ2p
π2p
σ*2s
σ2s
Energy
Bond energy /kJmol⁻¹
900
600
300
150
100
50
Bond length/pm
290
159
620
131
945
110
498
121
159
143
Bond order
1
2
3
2
1
0
Magnetic properties
Paramagnetic
Diamagnetic
Diamagnetic
Paramagnetic
Diamagnetic
–
Valence electron configuration
(σ2s)2 (σ*2s)2 (π2p)2
(σ2s)2 (σ*2s)2 (π2p)4
(σ2s)2 (σ*2s)2 (π2p)4 (σ2p)2
(σ2s)2 (σ*2s)2 (σ2p)2 (π2p)4 (π*2p)2
(σ2s)2 (σ*2s)2 (σ2p)2 (π2p)4 (π*2p)4
(σ2s)2 (σ*2s)2 (σ2p)2 (π2p)4 (π*2p)4 (σ*2p)2

for upto 14 e^-

(B_2, C_2, N_2)



degenerate orbitals

S-P Mixing

$N_2 \Rightarrow 14e^-$

$$B.O. = \frac{N_B - N_{AB}}{2}$$

$$\frac{10-4}{2} = 3$$

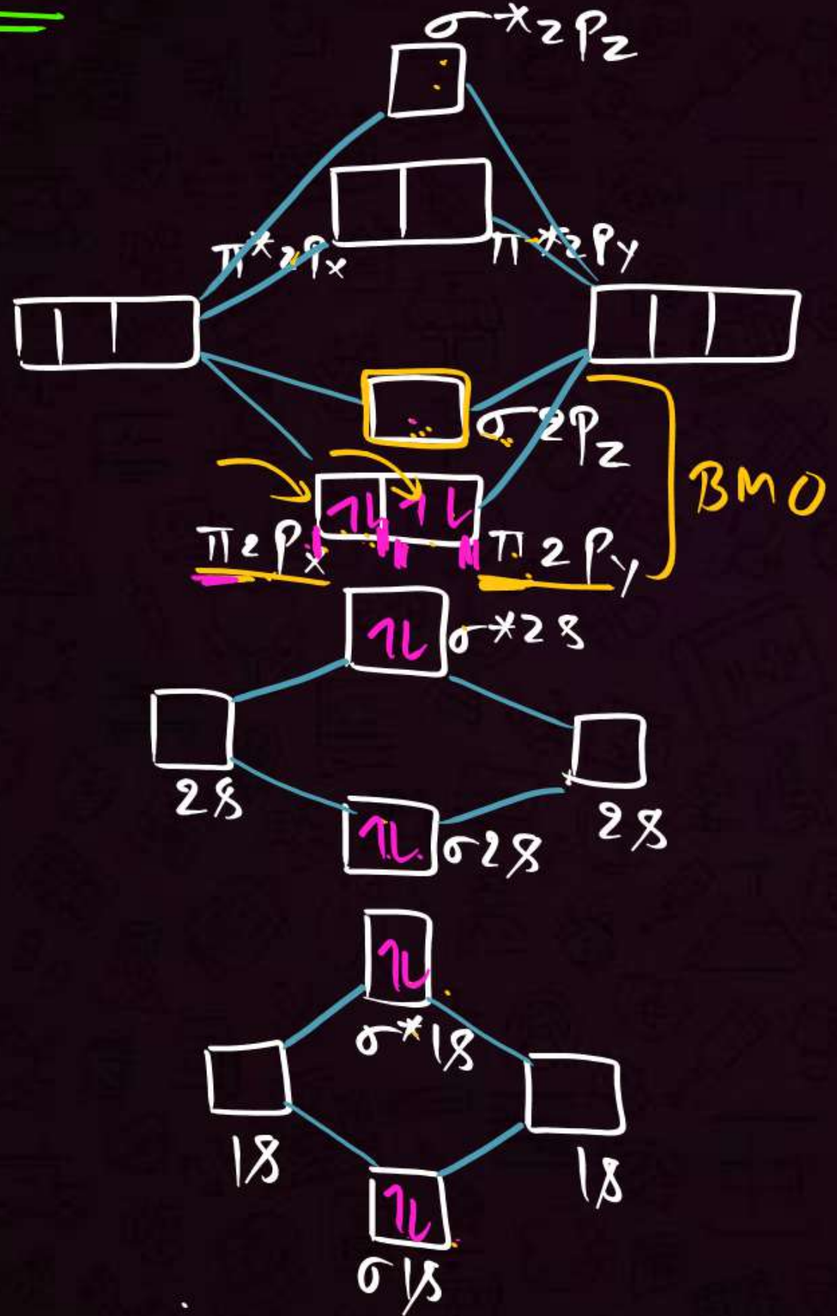
$N \equiv N$ (one σ and two π bonds)

for upto 14 e^-

(B_2, C_2, N_2)

$B_2 = 10e^-$

$C_2 = 12e^-$



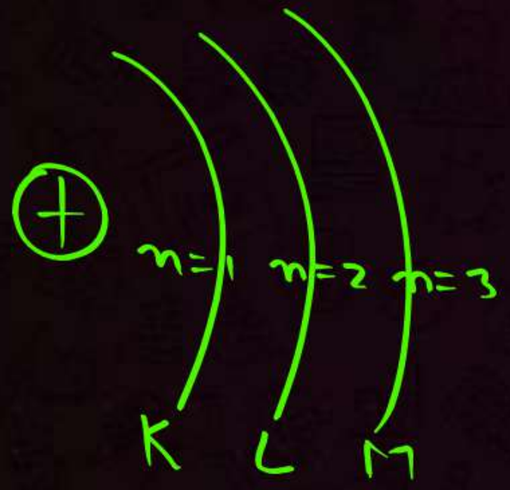
$B_2 \Rightarrow B.O. = \frac{6-4}{2} = 1$ (π-Bond)

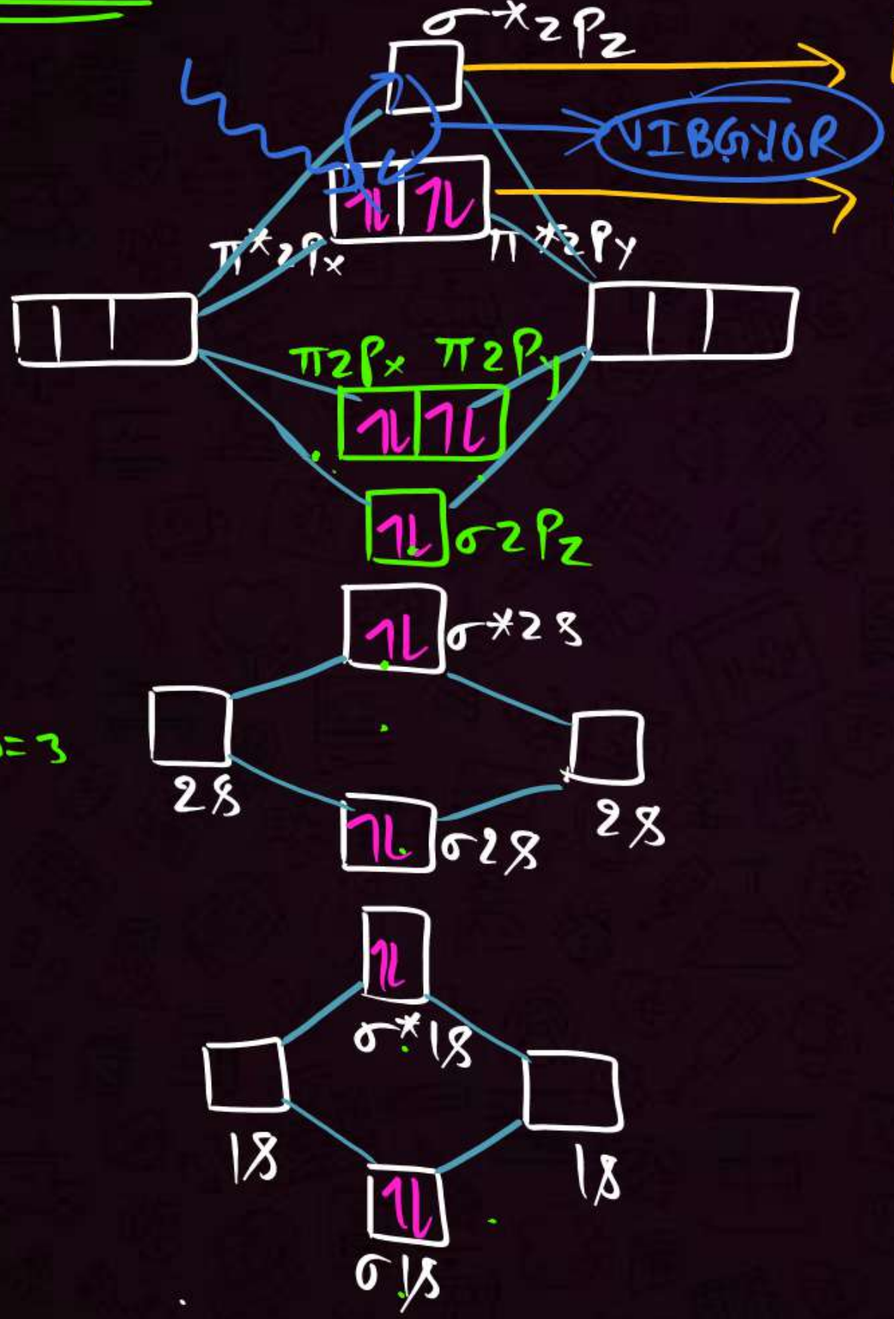
$C_2 \Rightarrow B.O = \frac{8-4}{2} = 2$ (π-Bond)

for More than 14e⁻

$(O_2, F_2,)$

18e⁻





LUMO (least Unoccupied M.O.)

HOMO (Highest occupied M.O.)

KK

$F_2 \rightarrow \sigma 1s^2, \sigma^* 1s^2, \sigma 2s^2, \sigma^* 2s^2, \sigma 2p_z^2$

$\pi 2p_x^2 = \pi 2p_y^2, \pi^* 2p_m^2 = \pi^* 2p_y^2$

Hitrick

- N_2 = 14e⁻ = 3.0, Dia
- B_2 = 10e⁻ = 1.0, para
- C_2 = 12e⁻ = 2.0 (Dia.)

- N_2^- = 15e⁻ = 2.5 (Para.)
- N_2^+ = 13e⁻ = 2.5 (para.)

20 → 0
19 → 0.5
18 → 1.0
17 → 1.5
16 → 2.0
15 → 2.5
14 → 3.0
13 → 2.5
12 → 2.0
11 → 1.5
10 → 1.0
9 → 0.5
8 → 0

- O_2 = 16e⁻ = 2.0 (para)
- O_2^+ = 15e⁻ = 2.5 (para)
- O_2^- = 17e⁻ = 1.5 (para)

odd No. of e⁻ ⇒ Para (1 unpaired e⁻)

even no. of e⁻ ⇒ Diamagnetic

(Except ⇒ 10, 16, 32 e⁻ (Paramagnetic 2 unpaired e⁻))

→ Halogens are coloured due to Homo lumo transition

F_2 → yellow

Cl_2 → Greenish Yellow

Br_2 → Red / Reddish Brown

I_2 → Violet

QUESTION

In which of the following processes, the bond order has increased and paramagnetic character has changed to diamagnetic?

(2019 Main, 9 Jan II)

(A) $O_2 \rightarrow O_2^+$

(B) $N_2 \rightarrow N_2^+$

(C) $O_2 \rightarrow O_2^{2-}$

(D) $NO \rightarrow NO^+$

QUESTION

The species having bond order different from that in CO is (2007, 3M)

$14e^- \Rightarrow 3.0$

A NO^- $16e^- = 2.0$

B NO^+ $14e^- = 3.0$

C CN^- $14e^- = 3.0$

D N_2 3.0

QUESTION

Among the following, the paramagnetic compound is (2007, 3M)

(A) Na_2O_2 → $2Na^+ + O_2^{-2}$ 18e⁻ = Dia

(B) O_3

(C) N_2O

$N \equiv N \rightarrow O$

(D) KO_2 → $K^+ + O_2^-$ 17e⁻ ⇒ para.

QUESTION

Which of the following molecular species has unpaired electron (s)? (2002, 3M)

A N_2

B F_2

C O_2^-

D O_2^{2-}

QUESTION

The common features among the species CN^-, CO and NO^+ are (2001, 1M)

(A) bond order three and isoelectronic

(B) bond order three and weak field ligands

(C) bond order two and acceptors

(D) isoelectronic and weak field ligands

* Note :→

① $NO \Rightarrow 15e^- = 2.5$

$NO^+ \Rightarrow 14e^- = 3.0$

② $CO \Rightarrow 14e^- = 3.0$

$CO^+ \Rightarrow \underline{13e^-} \Rightarrow$ ~~2.5~~ $\underline{> 3.0}$

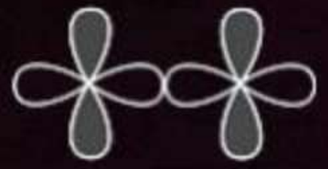
* Reason :→ e^- removed from CO to form CO^+ was present in N.B.M.O having slightly A.B characters. So B.O. ↑se slightly.

QUESTION

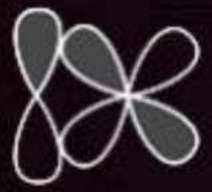
H.W.

Match the orbital overlap figures shown in Column I with the description given in Column II and select the correct answer using the codes given below the Columns.

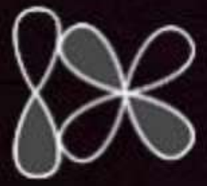
Column I **Column II** **(2014 Adv.)**

A. 

1. p-d π antibonding

B. 

2. d-d σ bonding

C. 

3. p-d π bonding

D. 

4. d-d σ antibonding

Codes:

	A	B	C	D
A	4	3	2	1
B	1	2	3	4
C	2	3	1	4
D	4	1	2	3

QUESTION

Two pi and half sigma bonds are present in

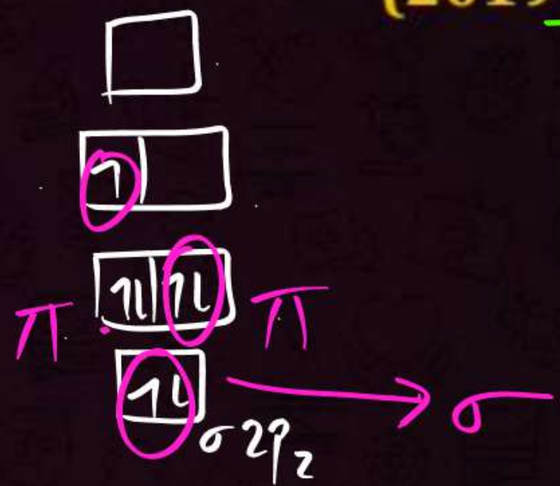
(2019 Main, 10 Jan 1)

A O_2^+

B N_2

C N_2^+

D O_2



1σ 1π 1/2 π

QUESTION

Which of the following compounds contain(s) no covalent bond(s)?

KCl, PH_3, O_2, B_2H_6, H_2SO_4, CsF

(2018 Main)

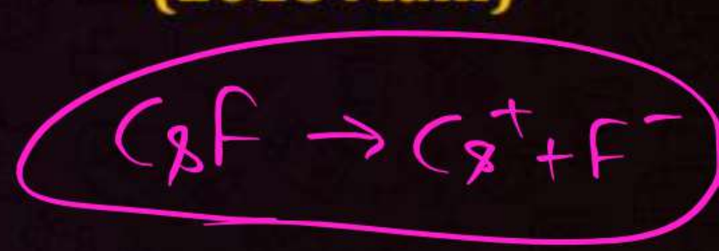
K^+, Cl^-

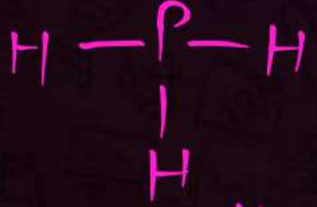
A. KCl, B_2H_6, PH_3

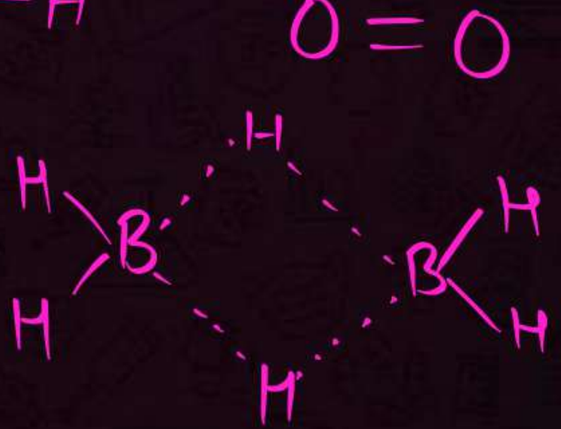
B. KCl, H_2SO_4

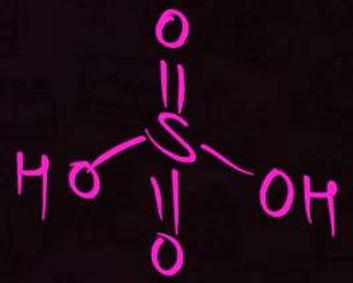
C. KCl, CsF

D. KCl, B_2H_6





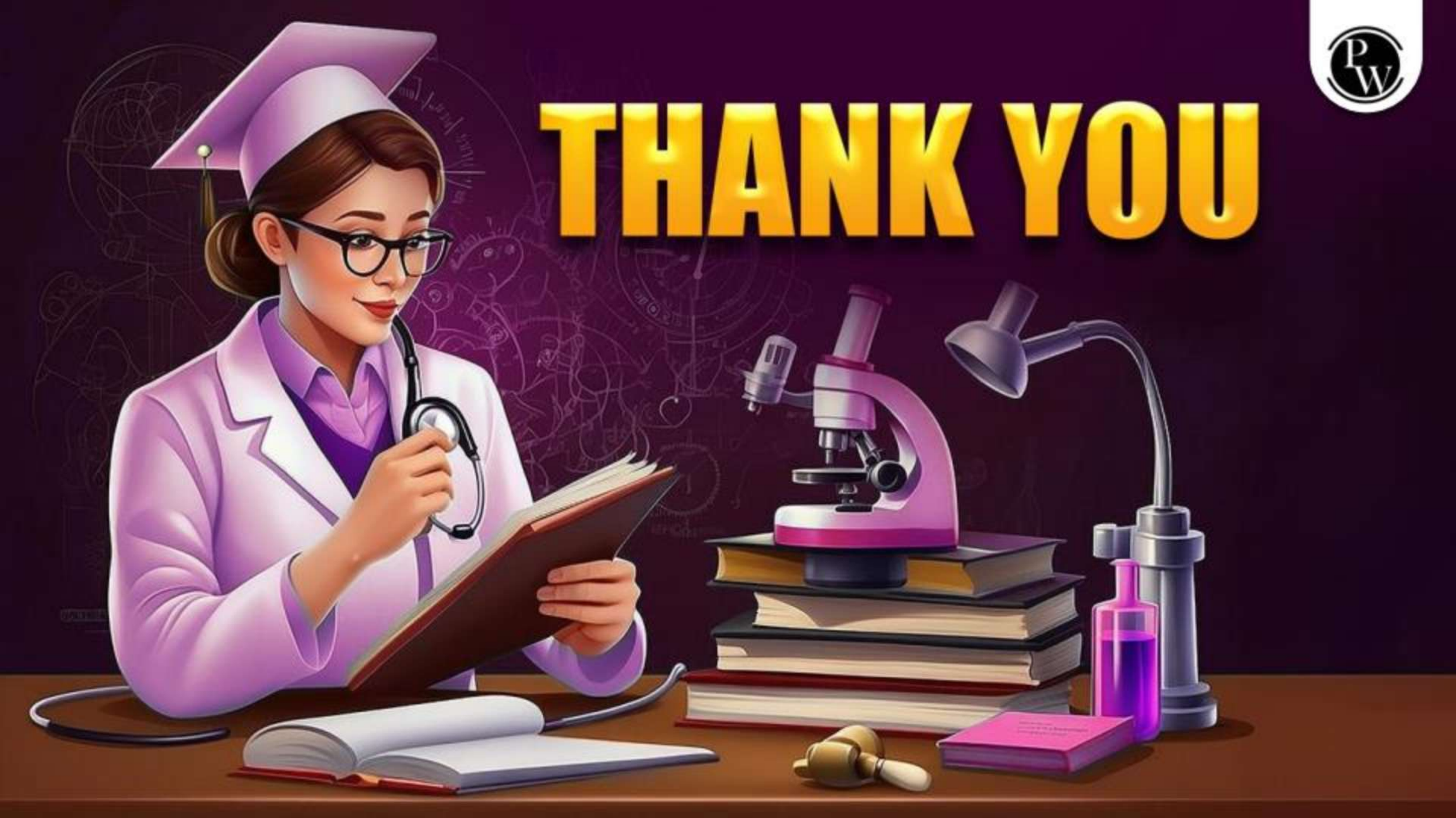




Homework

Practice Sheet

FOR NOTES & DPP CHECK DESCRIPTION


THANK YOU