Chemistry - 2

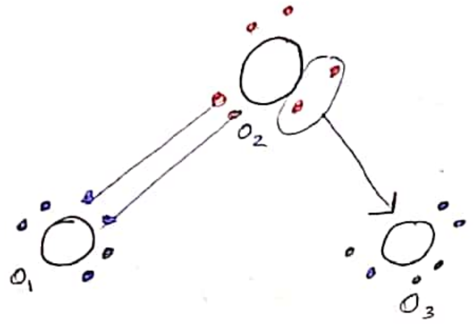
Ch - Chemical Bonding

Formal Charge :-

→ It is defined for Lewis dot structures.
→ It is the charge on atom in a given molecule or ion.

{F·C} = {no. of valence e^-} − {no. of bonds formed by that atom} − {total no. of e^- of l·P}

Eg ①



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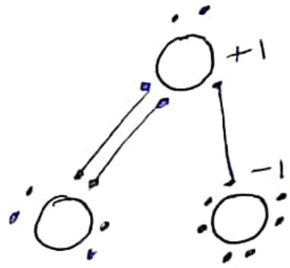
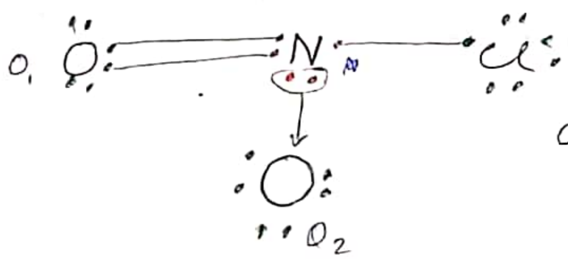


$F \cdot C (O_1) = 6 - 2 - 4 = 0$

$F \cdot C \cdot (O_2) = 6 - 3 - 2 = +1$

$F \cdot C \cdot (O_3) = 6 - 1 - 6 = -1$

② NO_2Cl



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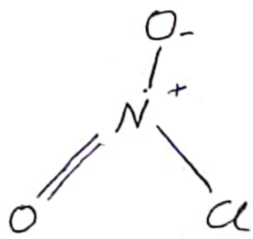


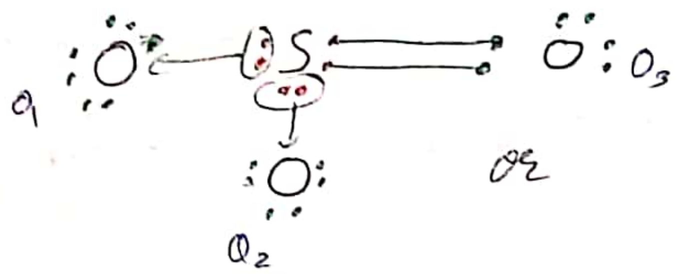
$F \cdot C (O_1) = 6 - 2 - 4 = 0$

$F \cdot C (O_2) = 6 - 1 - 6 = -1$

$F \cdot C (N) = 5 - 4 - 0 = +1$

$F \cdot C (Cl) = 7 - 1 - 6 = 0$

③ SO_3



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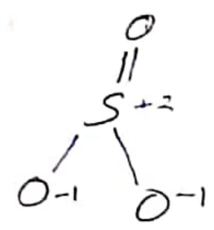


$F.C.(S) = 6 - 4 - 0 = +2$

$FC.(O_1) = 6 - 1 - 6 = -1$

$FC.(O_2) = 6 - 1 - 6 = -1$

$FC.(O_3) = 6 - 2 - 4 = 0$

Q. which of the following species do not follow octet rule?

① CCl_4

② NF_3

✓③ AlF_3

✓④ ClF_3

Q. Which of the following species do not follow octet rule?

① H_2Se → ✓

✓② NO_2 → odd e^-

✓③ BCl_3 → Hypovalent

✓④ XeO_4 → Hypervalent

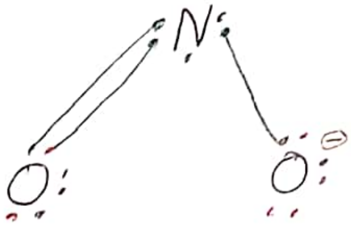
Note – ① All second period elements have lewis dot structures same as actual structures

② 3rd period elements have increased co-valency due to excitation so instead of forming co-ordinate bond they prefer to form more stable double bond.

③ Whenever a co-ordinate bond is formed by a 3rd period element in lewis dot structure, replace it with a double bond to get actual structure.

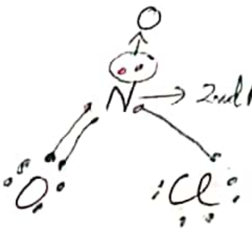
④ Structure having least formal charge on atom is most stable structure.

Ex ①. NO_2^-



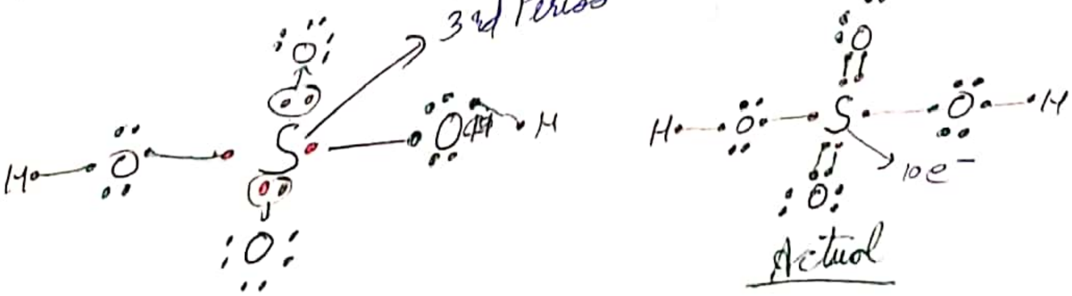
LDS ✓
Actual ✓

② NO_2Cl



LDS ✓
2nd Period
Actual ✓

③ H_2SO_4

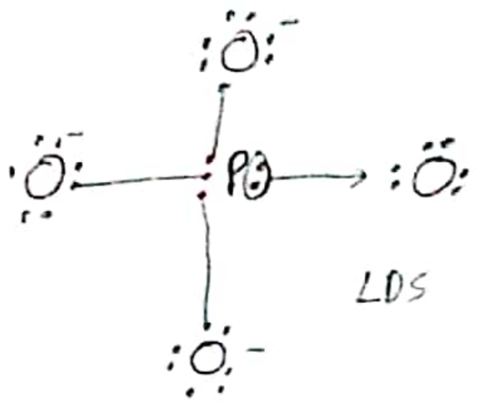


3rd Period

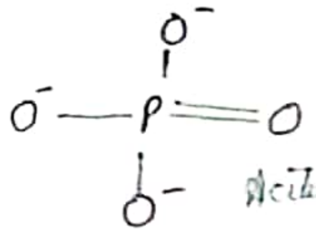
~~Actual~~
LDS

Actual

(4) PO_4^{3-}

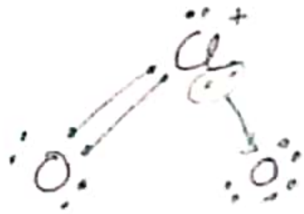


LDS

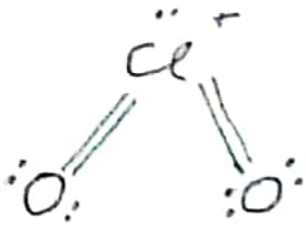


Actual/Covalent structure

(5) ClO_2^+



LDS

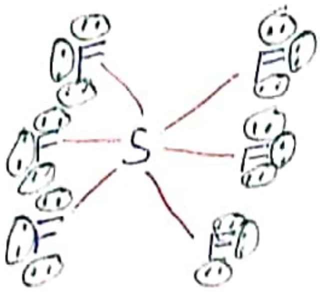


Actual

Q Calculate Total no. of L.P.

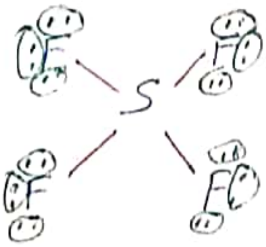
1. SF_6
2. SF_4
3. XeF_2
4. XeF_4
5. SO_2
6. SO_3
7. $NOCl_3$

(1) SF_6



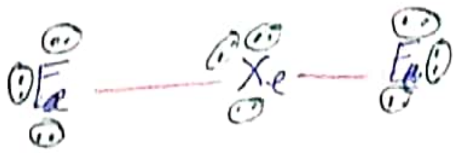
$L.P = 3 \times 6$
$= 18$

(2) SF_4



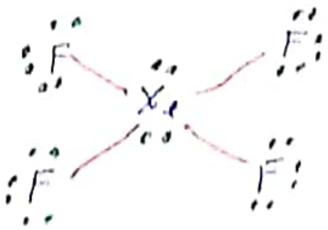
$L.P = 3 \times 4 + 1$
$= 12 + 1$
$= 13$

(3) XeF_2



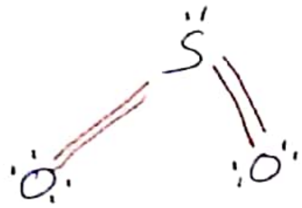
$L.P = 3 \times 2 + 3$
$= 6 + 3$
$= 9$

(4) XeF_4



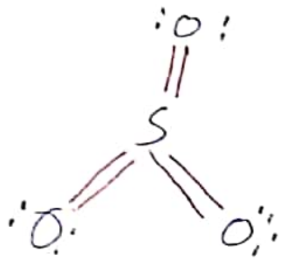
$L.P = 4 \times 3 + 2$
$= 12 + 2$
$= 14$

⑤ SO_2



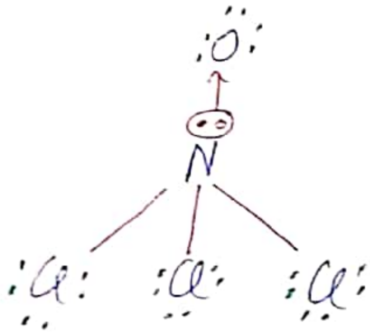
$L.P = 2 \times 2 + 1$
$= 4 + 1$
$= 5$

⑥ SO_3



$L.P = 2 \times 3$
$= 6$

⑦ $NOCl_3$



$L.P. = 3 + 3 \times 3$
$= 9 + 3$
$= 12$

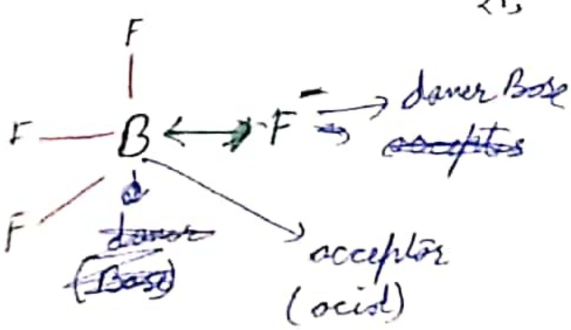
Note :-

① L.P. acceptor → ~~Acid~~ Lewis Acid

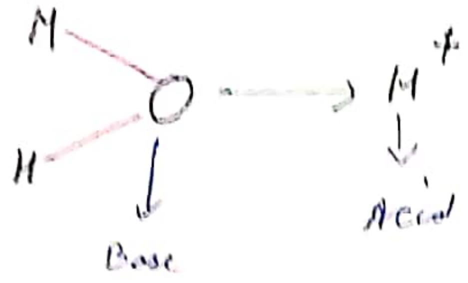
② L.P. Donor ⇒ ~~Bas~~ Lewis Base

Eg ① BF_4^-

B : $\boxed{\uparrow\downarrow}$ $1s^2$ $\boxed{\uparrow}$ $2s^2$ $\boxed{\uparrow \,|\, \uparrow \,|\,\;}$ $2p$ → vacant orbital



② H_3O^+



$H_2O + H^+$
(base) (acid)

Oxidation Number -

→ It is the hypothetical charge on atom when each bond is assumed to be ionic

→ Rules for calculating oxidation number.

① Oxidation no. of a free molecule or pure element is zero.
eg. Na, Cl_2, N_2, O_2, H_2, Ar, Zn

② Oxidation no. of alkali metals in compound state.
(Li, N, K, Rb, Cs, Fr) ⇒ ~~+2~~ +1
eg. LiCl, NaCl, Na_2SO_4

③ Oxidation no. of Alkaline earth metals in compound state -
(Be, Mg, Ca, Sr, Ba) ⇒ +2
eg. $Ca(OH)_2$, $MgCO_3$, $BeCO_3$, $Sr(OH)_2$, CaO

④ Oxidation no. of H :-
→ If ~~bond~~ bonded with metals → -1
→ If bonded with other element/ non- metals → +1
eg. $C\overset{+1}{H_4}$, $Li\overset{-1}{H}$, $\overset{+1}{H}NO_3$, $Ca\overset{-1}{H_2}$, $\overset{+1}{H_2}SO_4$

⑤ Oxidation no. of oxygen atom in compound. ⇒ −2 (In general)

: In peroxide (O_2^{2-}) ⇒ −1

in super oxide (O_2^{-1}) ⇒ −½

eg H_2O^{-2}, $H_2O_2^{-1}$, $KO_2^{-1/2}$, $O^{+2}F_2$, $O_2^{+1}F_2$

⑥ Oxidation Number of Florine in compound state ⇒ −1

⑦ Oxidation Number of Clorine, Bromine & Iodine as ~~a~~ surrounding atom ⇒ −1

as central atom. S can increase.

H.W.

Q ① SO_2

$S + 2(-2)$

$S = -4$

$\boxed{S = +4}$ ✓

② SO_3

③ $KMnO_4$

④ HNO_3

⑤ H_2O_2

⑥ H_2SO_4

⑦ KO_2

⑧ SO_4^{2-}

⑨ SO_3^{2-}

⑩ NH_4^+

⑪ CCl_4

⑫ MnO_4^{2-}

⑬ $Cr_2O_7^{2-}$

⑭ OF_2

⑮ VO

⑯ V_2O_5

⑰ $HClO_4$

⑱ N_2O_3

⑲ NO

⑳ $PH_4^{\oplus}$

㉑ $HClO_3$

㉒ $HClO_2$

㉓ $HClO$

㉔ CrO_4^{2-}

㉕ NO_2^-

㉖ AlN

㉗ Ca_3N_2

㉘ KNO_2

㉙ KNO_3

㉚ $Al_2(SO_4)_3$

㉛ NH_4NO_3

㉜ ZnO

㉝ Fe_2O_3

㉞ Pb_3O_4

㉟ Fe_3O_4

㊱ $Cr_2(SO_4)_3$

㊲ PO_4^{3-}

㊳ $H_2PO_3^{-1}$

㊴ HIO_4

㊵ H_3PO_4

㊶ H_3PO_2

㊷ CaH_2

㊸ ClO_3^-

㊹ BF_3

㊺ BeF_4^{2-}

㊻ $MgCO_3$

㊼ $NaHSO_4$

㊽ H_2S

㊾ N_3^-

㊿ Al

(51) $Na_2S_2O_3$

(2) SO_3

$S + 3(-2)$

$S + -6 = 0$

$\boxed{S = 6}$ ✓

(3) $Mn = 8 - 1$

$\boxed{Mn = 7}$ ✓

(3)(10) $2 + -8 + S = 0$

$S = 8 - 2$

$\boxed{S = 6}$ ✓

(4) $+1 - 6 + N = 0$

$N = 6 - 1$

$\boxed{N = 5}$ ✓

(5) $2 + 2O = 0$

$2O = 2$

$\boxed{O = -1}$ ✓

(6)(7) $1 + 2O = 0$

$2O = -1$

$\boxed{O = -1/2}$ ✓

(8) $S - 8 = -2$

$S = 8 - 2$

$\boxed{S = 6}$ ✓

(9) ~~S~~

~~S - 2 = 8 6~~

$\boxed{\cancel{S = 8}}$

$S = 6 - 2$

$\boxed{S = 4}$ ✓

(10) $N + 4 = 1$

$N = 1 - 4$

$\boxed{N = -3}$ ✓

(11) $\boxed{+4}$ ✓

(12) $Mn - 8 = -2$

$\boxed{Mn = 6}$ ✓

(13) $2Cr - 14 = -2$

$2Cr = 12$

$\boxed{Cr = 6}$ ✓

(14) $O - 2 = 0$

$\boxed{O = +2}$ ✓

(15) $V - 2 = 0$

$\boxed{V = 2}$ ✓

(16) $2V - 10 = 0$

$V = 10/2$

$\boxed{V = 5}$ ✓

(17) $1 + Cl - 8 = 0$

$\boxed{Cl = 7}$ ✓

(18) $2N - 6 = 0$

$\boxed{N = 3}$ ✓

(19) $N - 2 = 0$

$\boxed{N = 2}$ ✓

(20) $P + 4 = 1$

$\boxed{P = -3}$ ✓

(21) $1 + Cl - 6 = 0$

$\boxed{Cl = 5}$ ✓

(22) $1 + Cl - 4 = 0$

$\boxed{Cl = 3}$ ✓

(23) $\boxed{Cl = 1}$ ✓

Note -

① For any element range of O.S ⇒ N to N-8 (both included)

oxidation State

N ⇒ no of valence e^-

eg ① N: N=5

N: {+5 to -3}

② S: N=6

S: {+6 to -2}

③ O: N=6

O: {-2 to +6} (Expected)

O: {-2 to +2} (Actual) ⇒ Exception

④ Cl: N=7

Cl: {-1 to +7}

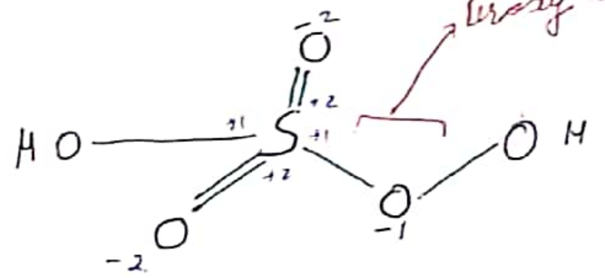
Eg H_2SO_5 (caro's acid)

$2 + x - 10 = 0$

$x = 10 - 2$

$\boxed{x = 8}$ → but had only till (-2 to 6)

so wrong



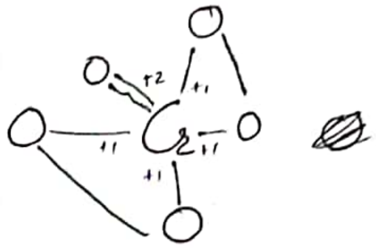
$2+2+1+1 = 6$

$\boxed{O.S\ of\ S = 6}$

Eg. CrO_5

$x - 10 = 0$

$x = 10$ ✗ but chromium has only 6 valence e^-



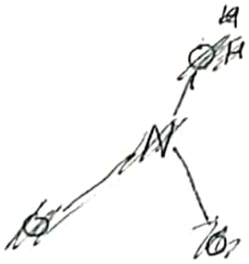
butterfly structure

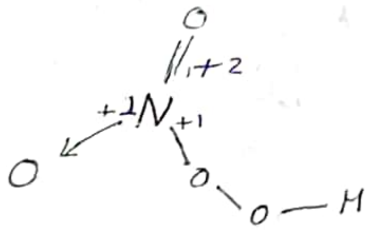
$2+1+1+1+1+1 = \boxed{6}$

Eg HNO_4

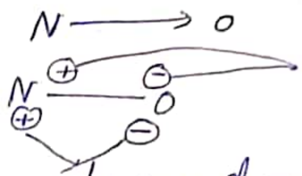
$N = 8-1$

$= 7$ ✗





$2+2+1 = \boxed{5}$



cause of co-ordinate bond

because of covalent bond

Note- Covalency depends on no. of bonds but oxidation state does not.

- May or may not equal

Covalency	Oxidation state
(1) No. of unpaired e^- in ground state or excited state.	(1) Permanent formal charge.
(2) Can't be negative or fractional	(2) Can be $\ominus$ve or fractional.
(3) Depends on the no. of bonds	(3) Do not depend on no. of bonds.

Q find covalency and oxidation state of

(1) NH_3 & NH_4^+

NH_3 oxidation state (N) @ — covalency (N)

$\boxed{1|1|1}$

$N - 3 = 0$

$\boxed{N = 3}$ — $\boxed{C.V. = 3}$

~~(2) BF_3 & BF~~

NH_4^+

~~$N = 1 + 4$~~

$\boxed{\cancel{N = 5}}$ — $\boxed{1|1|1|1}$ — $\boxed{C.V = 4}$

$N = 1 - 4$

$\boxed{= -3}$

(2) BF_3 & BF_4^-

$O.S = 3$ — $O.S = 3$

$C.V = 3$ — $C.V = 4$

③ XeO_2F_2 — $O.S = 6$, $C.V = 6$

XeO_3 — $O.S = 6$, $C.V = 6$

XeO_4 — $O.S = 8$, $C.V = 8$

Bond Order (B.O):-

→ It is the effective no. of bonds formed between ~~a~~ two atoms.

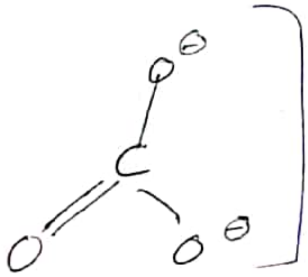
H_2 $H-H$ $B.O = 1$

O_2 $O=O$ $B.O = 2$

N_2 $N\equiv N$ $B.O = 3$

HCN $H-C\equiv N$ $B.O(C-H) = 1$, $B.O(C-N) = 3$

CO_3^{2-}



need resonance concept to calculate.

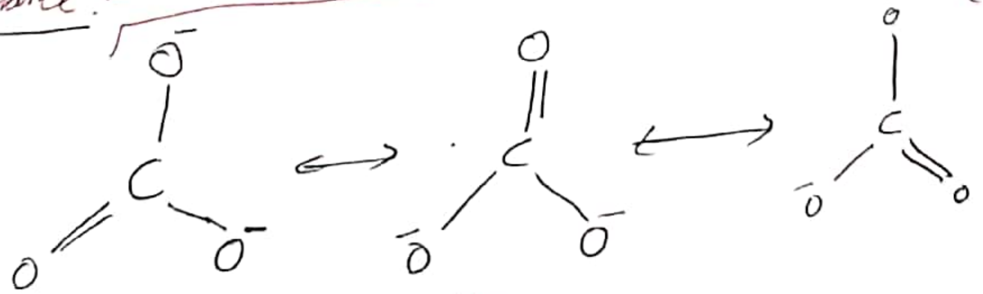
CO_2 $O=C=O$ ~~B.O (2)~~

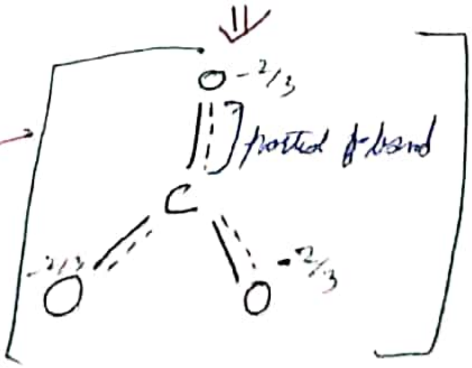
$B.O.(\cancel{=}) = 2$

resonating structures or canonical

Resonance:-

CO_3^{2-}





more stable

⇒ Resonance Hybrid.

$B.O = 1 + \frac{1}{3} = \frac{4}{3}$ $\boxed{= 1.333}$

$$\boxed{B.O = \frac{\sigma + \pi}{\sigma}}$$

→ σ is the first bond formed between two atoms

→ π is the second bond formed between two atoms.

Resonance - when a molecule cannot be completely represented by a single structure, It means a single structure cannot explain all the properties of a molecule Then it is said that the molecule is showing resonance.

→ True structure is said to ~~be~~ be '<u>Resonance Hybrid of the structure</u>'

→ This Phenomenon is called resonance and depends contributing structures is called <u>resonating or canonical structures.</u>

→ Resonance stabilize the molecule as the energy of Resonance hybrid is less than energy of any single canonicle structure.

<u>conditions for resonance</u>

(1) $\ominus$ve — $\overset{\to \text{ double bond}}{=}$

(2) L·P — =

(3) = — =

(4) $\oplus$ve — =

H. W.

24 $\boxed{G = 6}$ ✓

25 $\boxed{1}$ +3

26 $\boxed{-3}$ ✓

27 $\boxed{-3}$ ✓

28 $\boxed{3}$ ✓

29 $\boxed{5}$ ✓

30 $\boxed{6}$ ✓

31 $\boxed{1}$ -3, +5

32 $\boxed{2}$ ✓

33 $\boxed{3}$ ✓

34 $\boxed{8/3}$ ✓

35 $\boxed{+8/3}$ ✓

36 $\boxed{1, 6}$ +3, 6

37 $\boxed{5}$ ✓

38 $\boxed{3}$ ✓

39 $\boxed{7}$ ✓

40 $\boxed{5}$ ✓

41 $\boxed{1}$ ✓

42 $\boxed{-1}$ ✓

43 $\boxed{5}$ ✓

44 $\boxed{3}$ ~~2~~ ✓

45 $\boxed{2}$ ✓

46 $\boxed{2}$ ✓

47 $\boxed{6}$ ✓

48 $\boxed{-2}$ ✓

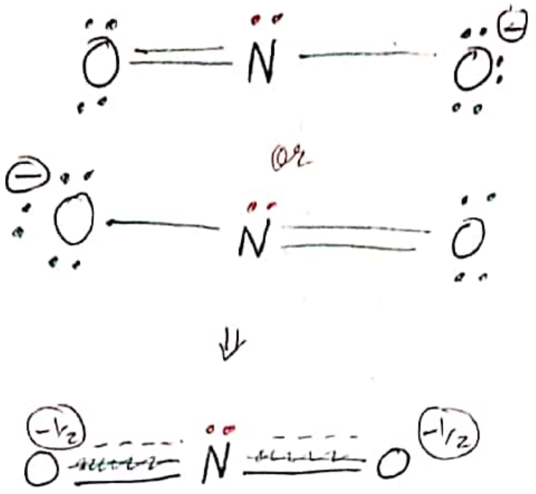
49 $\boxed{-1/3}$ ✓

50 $\boxed{0}$ ✓

51 $\boxed{2}$ ✓

31

Eg ① NO_2^-

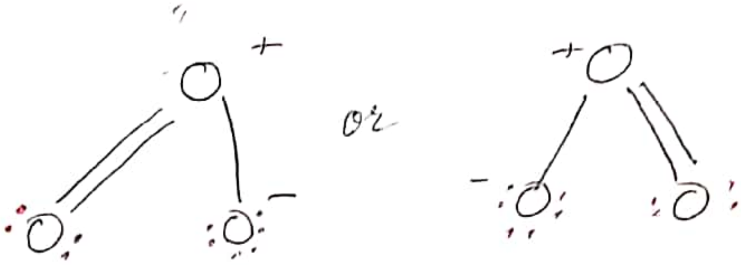


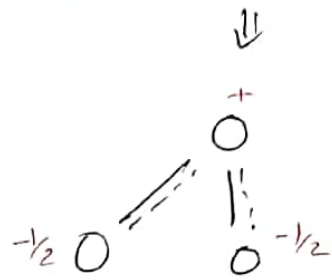
$B.O = 1 + \frac{1}{2}$

$= \frac{2+1}{2}$

$\boxed{= \frac{3}{2}}$

Eg ② O_3

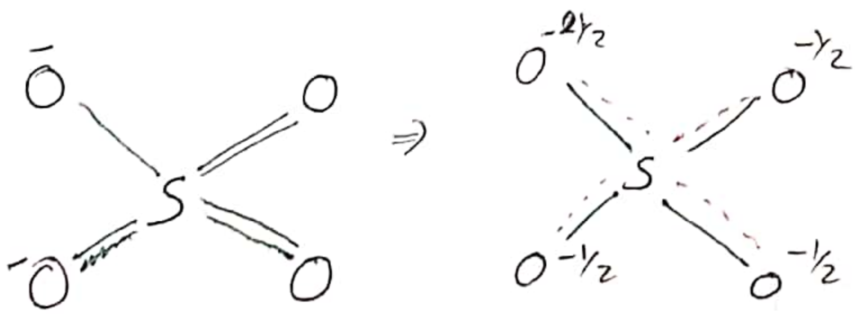




$B.O = 1 + \frac{1}{2}$

$\boxed{= 3/2}$

Eg ③ SO_4^{2-}

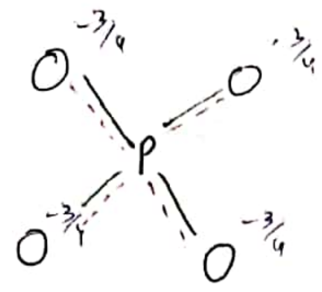


$B.O = \frac{4+2}{4}$

$= \frac{6}{4}$

$\boxed{= \frac{3}{2}}$

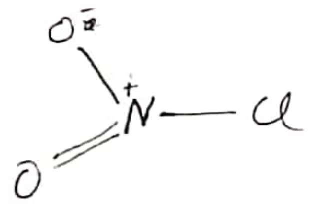
Eg (4) PO_4^{3-} ~~B.O~~



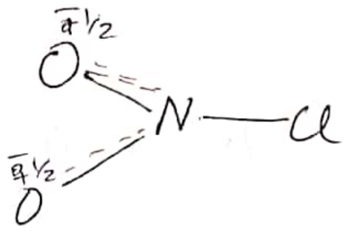
$B \cdot O = \frac{4+1}{4}$

$\boxed{= \frac{5}{4}}$

Eg (5). NO_2Cl



or

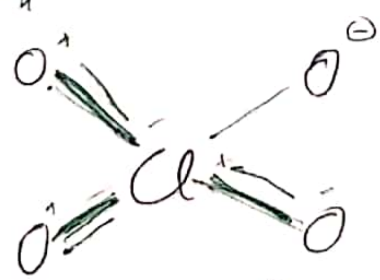


$B \cdot O\ (N-Cl) = \boxed{1}$

$B \cdot O\ (O-N) = \frac{2+1}{2}$

$\boxed{= \frac{3}{2}}$

Eg (6). ClO_4^-



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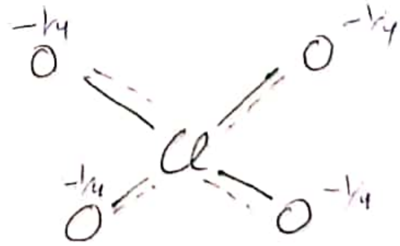


$B \cdot O = \frac{\cancel{11}^{\,9} + 3}{4}$

$\boxed{= \frac{7}{4}}$

Q Draw the Resonating Structure Of Azide ion ? (N_3^-)

$$N \equiv N^{\oplus} - {}^{\ominus}N^{\ominus}$$

B.O =

$$\overset{-2}{N} - \overset{\oplus}{N} \equiv N$$

$$\boxed{\overset{-}{N} = \overset{+}{N} = \overset{-}{N}}$$

$$B.O = \frac{4}{2} = \boxed{2}$$

$$\overset{-1}{N} \overset{\cdot\cdot\cdot}{=} \overset{+1}{N} \overset{\cdot\cdot\cdot}{=} \overset{-1}{N}$$ [Resonance Hybrid]

Note :-

(1) $\boxed{B.O \propto B.S \propto B.E \propto B.D.E \propto \frac{1}{B.L}}$

B.O → Bond Order
B.S → Bond Strength
B.E → Bond Energy
B.D.E → Bond Disossociaon energy
B.L → Bond Length

Q Compare bond length in (C-O)

CO CO_2 CO_3^{2-}

$B.O = 3 < B.O = 2 < B.O = \frac{4}{3} = 1.33$ (B.L)

$CO > CO_2 > CO_3^{2-}$ (B.E)

~~B.O~~ ~~CO~~

Q Compare (S-O) Bond Length in

SO_4^{2-} SO_3^{2-}

$B.O = 1.5 > B.O = \frac{4}{3} = 1.33$

$SO_4^{2-} < SO_3^{2-}$

Q Compare (Cl-O) Bond strength & Length

ClO_4^- ClO_3^-

$B.O = \frac{7}{4} = 1.75 > B.O = \frac{5}{3} = 1.66$

$ClO_4^- > ClO_3^-$ (B.S)

$ClO_4^- < ClO_3^-$ (B.L)

Q (N-O) bond length

NO_3^- NO_2^- NO_2^+

$B.O = \frac{4}{3} = 1.33$ $B.O = \frac{3}{2} = 1.5$ $B.O = 2$

$NO_3^- > NO_2^- > NO_2^+$ (B.L)

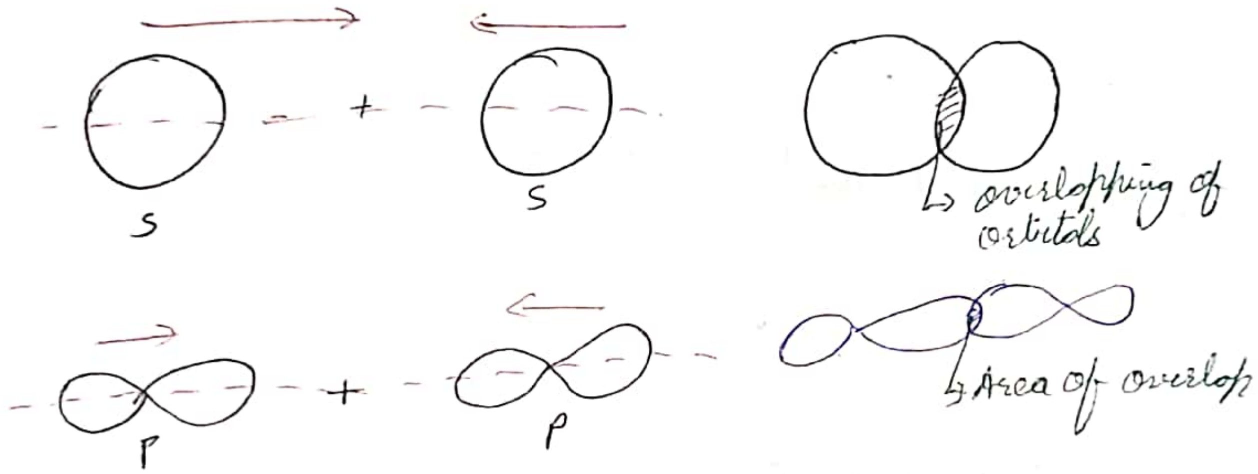
Valence Bond Theory (VBT)

→ It Explains how covalent bond will be formed.

→ Structures ~~were~~ are based on covalency.

→ It is based on overlapping of atomic orbitals.

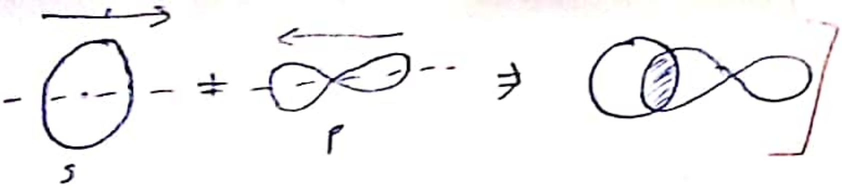
→ A bond is formed by overlapping ~~of~~ of pure atomic orbitals in a specified orientation.



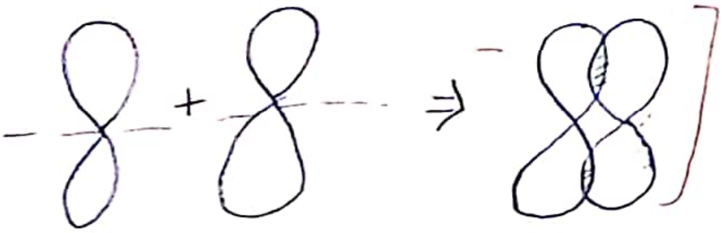
Note - If overlapping area increases then covalent strength increases

Overlapping

- Head on overlapping (co-axial overlapping) ⇓ σ - bond
- Sideways Overlapping (co-lateral overlapping) ⇓ π - bond



Co axial overlapping
σ - bond



Sideways overlapping
π - bond

Note → σ bond is stronger than π bond as overlapping area is more, electron density is more.

Covalent Bond

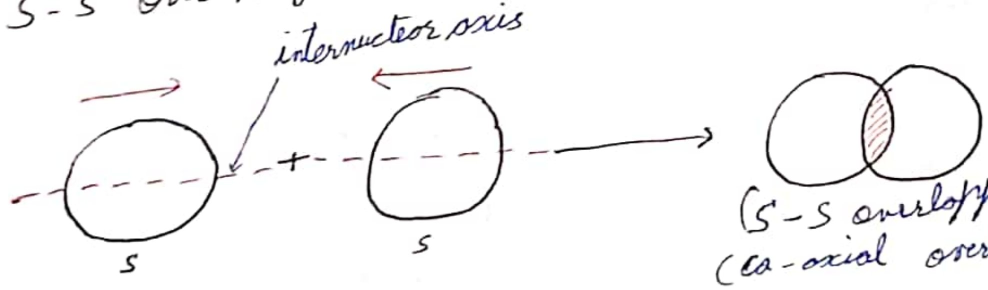
σ - bond.

1. s - s overlapping
2. s - p overlapping
3. p - p overlapping

π - bond

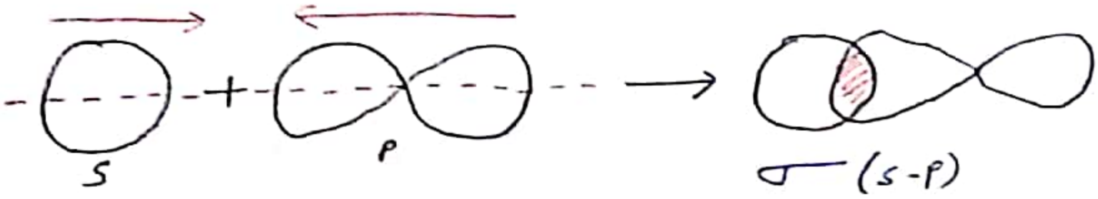
1. p - p orbital overlapping
2. p - d overlapping
3. d - d overlapping.

1. s - s overlapping

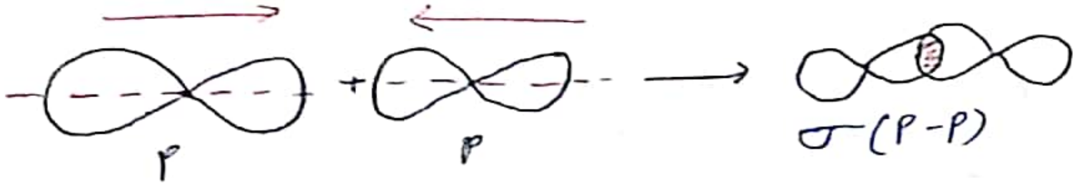


(s - s overlapping)
(co-axial overlapping)
$\sigma_{(s-s)}$

(2) S - P overlapping

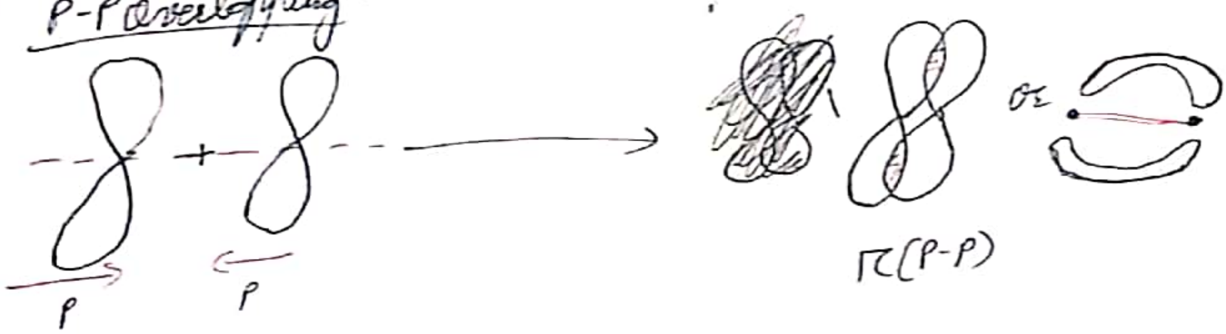


(3) p-p overlapping

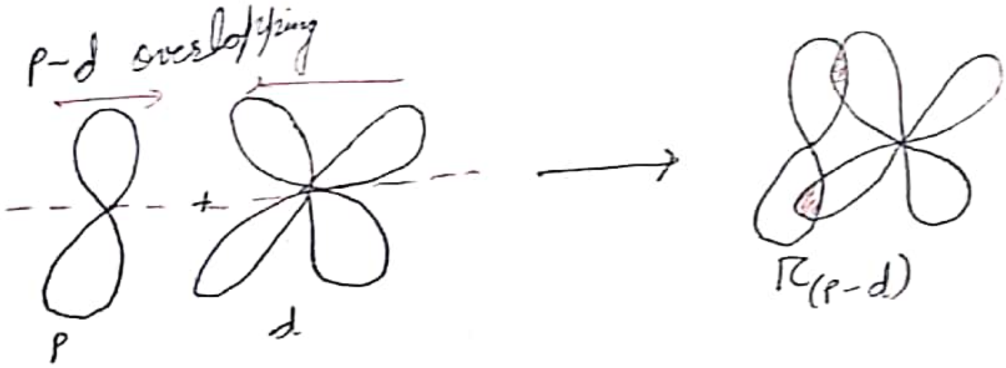


π- bonds

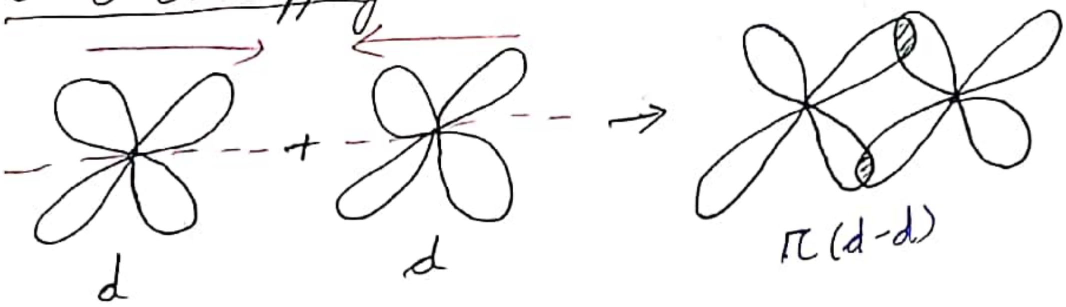
(4) p-p overlapping



(5) p-d overlapping

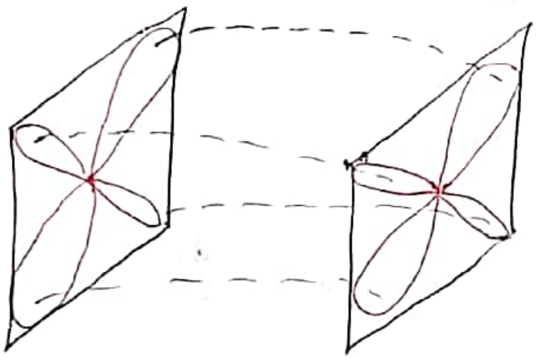


⑥ d - d overlapping



Delta Bond (δ):-
(Type of π bond)

→ It is a special type of π bond in which 4 lobes are interacted known as delta bond. (δ).



Note:- w.r.t one orbital

1. lobes interacted
2. lobes interacted
4. lobes interacted

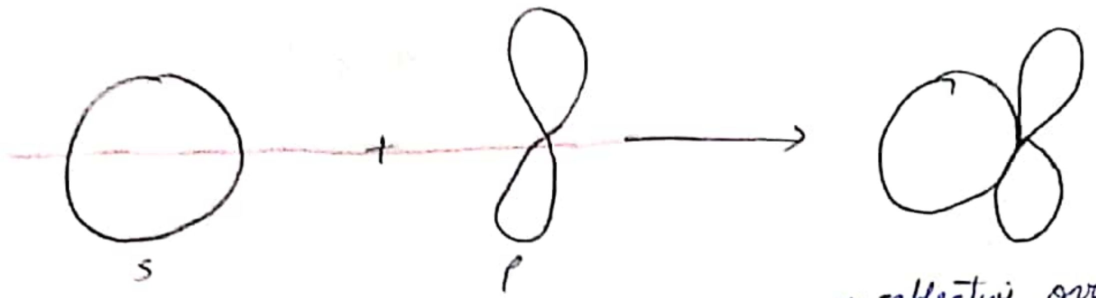
Types of bond

σ bond

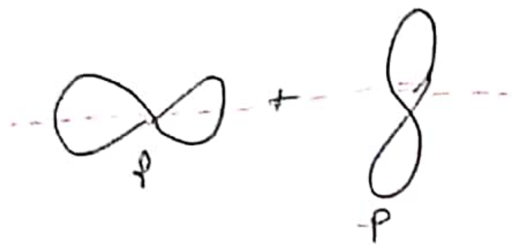
π bond

δ bond.

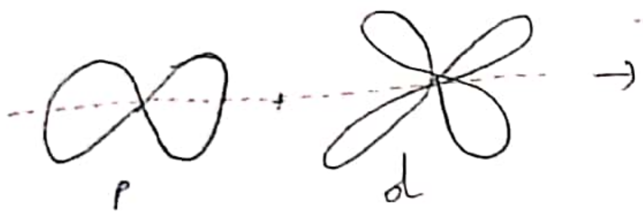
Zero Overlapping: Overlapping which are not acceptable because of ineffective overlapping.



no effective overlapping
∴ no bond.



→ no bond formation



→ no. bond formation

Note:-

1. s-orbital can never form π bond.
2. According to VBT, D-orbitals never form σ bond.
3. The first bond formed between two atoms is always a σ bond.
4. Once a ~~σ~~ σ bond is formed, the rest bonds formed after that will be π-bond.

Atomic orbital	Internuclear axis	Type of bond
(1) $s + p_x$	$x \longrightarrow$	σ
	$y \longrightarrow$	×
	$z \longrightarrow$	×
(2) $s + p_y$	$x \longrightarrow$	×
	$y \longrightarrow$	σ
	$z \longrightarrow$	×
(3) $s + p_z$	$x \longrightarrow$	×
	$y \longrightarrow$	×
	$z \longrightarrow$	σ
(4) $p_x + p_x$	$x \longrightarrow$	σ
	$y \longrightarrow$	π
	$z \longrightarrow$	π
(5) $p_y + p_y$	$x \longrightarrow$	π
	$y \longrightarrow$	σ
	$z \longrightarrow$	π
(6) $p_z + p_z$	$x \longrightarrow$	π
	$y \longrightarrow$	π
	$z \longrightarrow$	σ
(7) $p_x + p_y$	$x \longrightarrow$	×
	$y \longrightarrow$	×
	$z \longrightarrow$	×

8. $P_y + P_z$	$x \longrightarrow$ $y \longrightarrow$ $z \longrightarrow$	$\times$ $\times$ $\times$
9. $P_x + d_{xy}$	$x \longrightarrow$ $y \longrightarrow$ $z \longrightarrow$	$\times$ π $\times$
10. $P_y + d_{yz}$	$x \longrightarrow$ $y \longrightarrow$ $z \longrightarrow$	$\times$ $\times$ π
11. $d_{xy} + d_{xy}$	$x \longrightarrow$ $y \longrightarrow$ $z \longrightarrow$	π π δ
12. $d_{yz} + d_{yz}$	$x \longrightarrow$ $y \longrightarrow$ $z \longrightarrow$	δ π π
13. $d_{zx} + d_{zx}$	$x \longrightarrow$ $y \longrightarrow$ $z \longrightarrow$	π δ π
14. $d_{xy} + d_{yz}$	$x \longrightarrow$ $y \longrightarrow$ $z \longrightarrow$	$\times$ $\times$ $\times$
15. $d_{yz} + d_{zx}$	$x \longrightarrow$ $y \longrightarrow$ $z \longrightarrow$	$\times$ $\times$ $\times$
16. $d_{xy} + d_x d_z$	$x \longrightarrow$ $y \longrightarrow$ $z \longrightarrow$	$\times$ $\times$ $\times$

H.W. (24-05-2024)

O-1 $[1,7] \cup \{9\}$

O-2 $[1,5] \cup \{11\}$

S-1 $\{16\}$

S-2 $\{3, 4, 17, 18, 19\}$

Note -

1. Two Perpendicular orbitals can never form any bond
2. π bond is formed only after ~~sigma~~ σ bond.
3.

Q Calculate π bond & σ bond ?

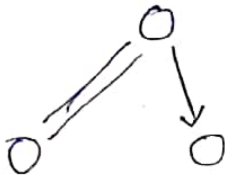
1. CO_2

$$O=C=O$$

π bonds - 2

σ bonds - 2

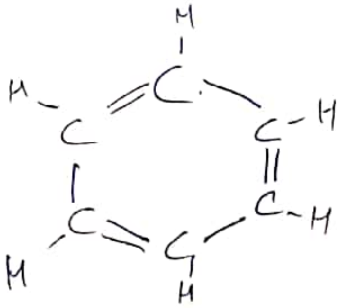
2. O_3



$\sigma = 2$

$\pi = 1$

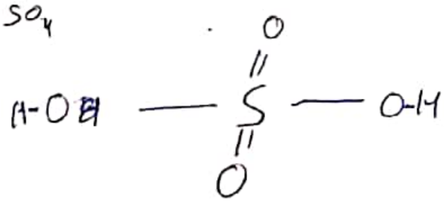
3.



$\sigma = 12$

$\pi = 3$

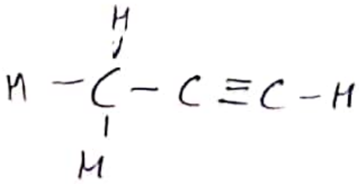
4. H_2SO_4



$\sigma = 6$

$\pi = 2$

5. $H_3C-C\equiv CH$



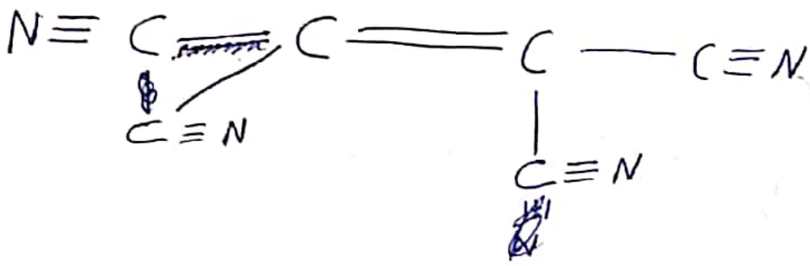
$\sigma = 6$

$\pi = 2$

⑥ HCN

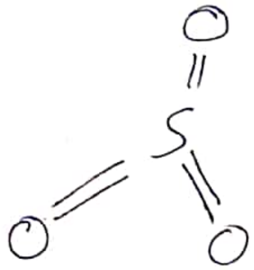
$H-C\equiv N$ $\qquad \sigma = 2$, $\pi = 2$

⑦ $C_2(CN)_4$



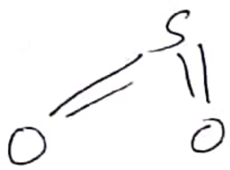
$\sigma = 9$
$\pi = 9$
$\therefore P = 4$
$B.P = 18\ (9+9)$

⑧ SO_3



$\sigma = 3$
$\pi = 3$

⑨ SO_2



$\sigma = 2$
$\pi = 2$

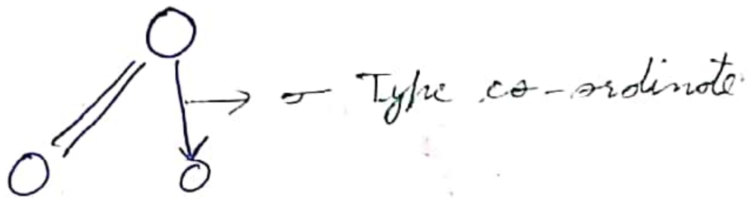
Note - ① Maximum 3 covalent bonds ($1\sigma + 2\pi$) can be formed between 2 atoms

Eg N_2: $N\equiv N$ $\qquad$ CO: $\overset{\ominus}{C}\equiv\overset{\oplus}{O}$ $\qquad$ C_2: ~~C≡C~~

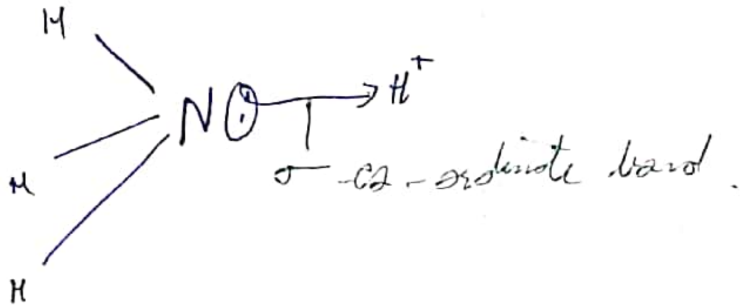
Types of Co-ordinate bonds

① σ Type-coordinate — If first bond is coordinate bond between co-ordinate atoms then it is known as σ Type co-ordinate bond

eg. O_3

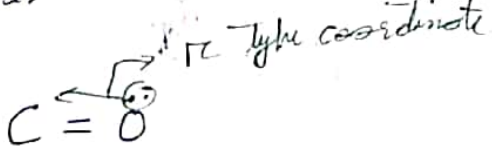


eg②.

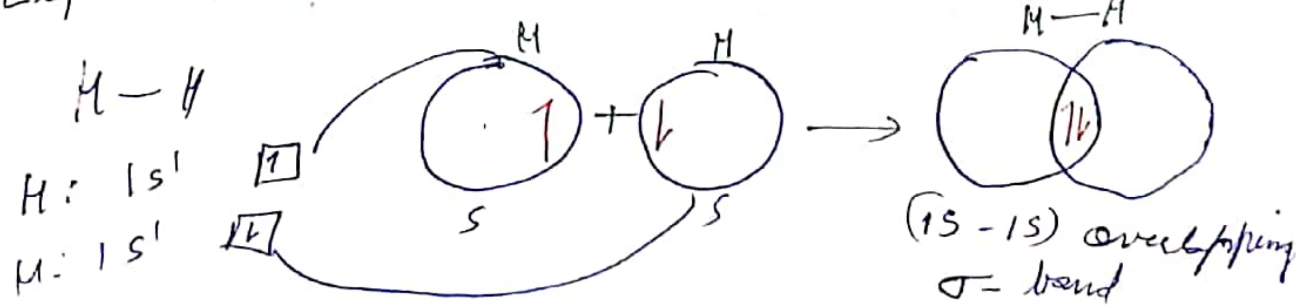


② π Type coordinate bond — If co-ordinate bond formed between 2 atoms is not first then it is known as π type coordinate bond.

eg①.



Q Explain H_2 molecule formation by VBT.

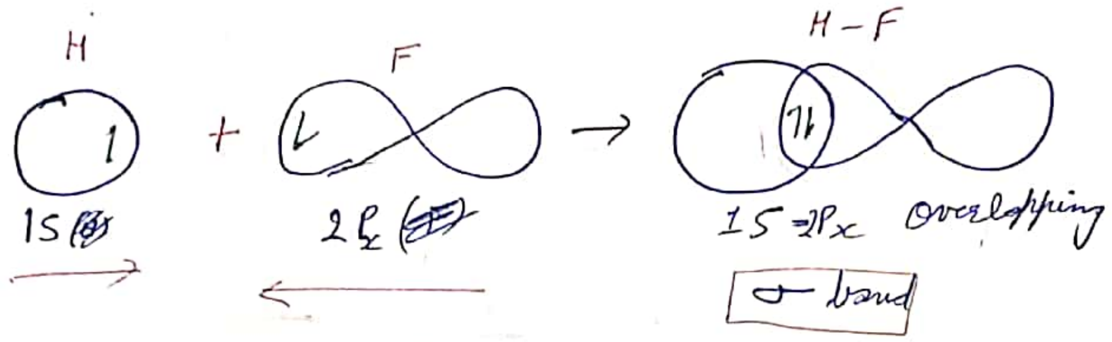


Q formation of HF explaining VBT

$H - \ddot{\underset{\cdot\cdot}{F}}:$

H: $1s^1$

F: $1s^2 2s^2 2p^5$

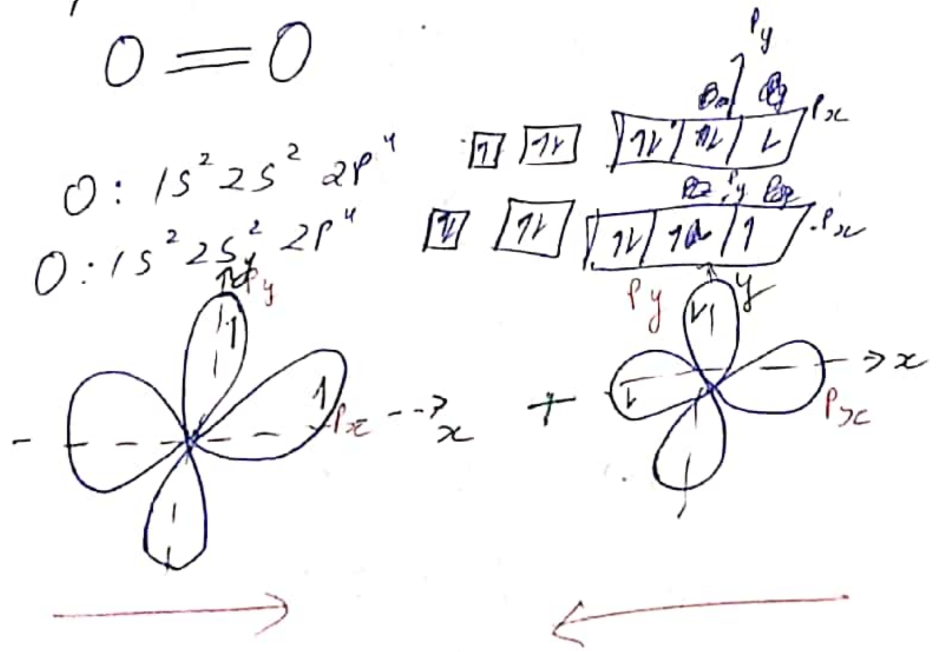


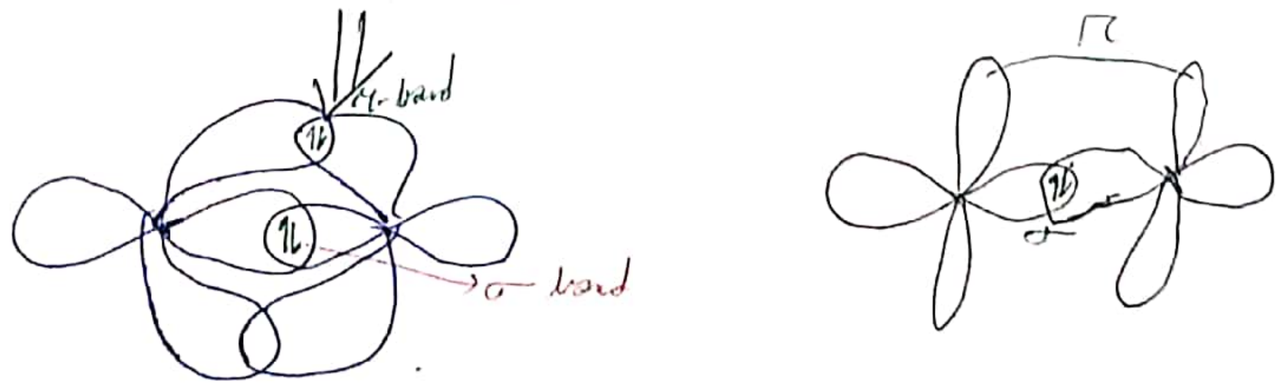
Q Explain O_2 formation using VBT.

$O = O$

O: $1s^2 2s^2 2p^4$

O: $1s^2 2s^2 2p^4$



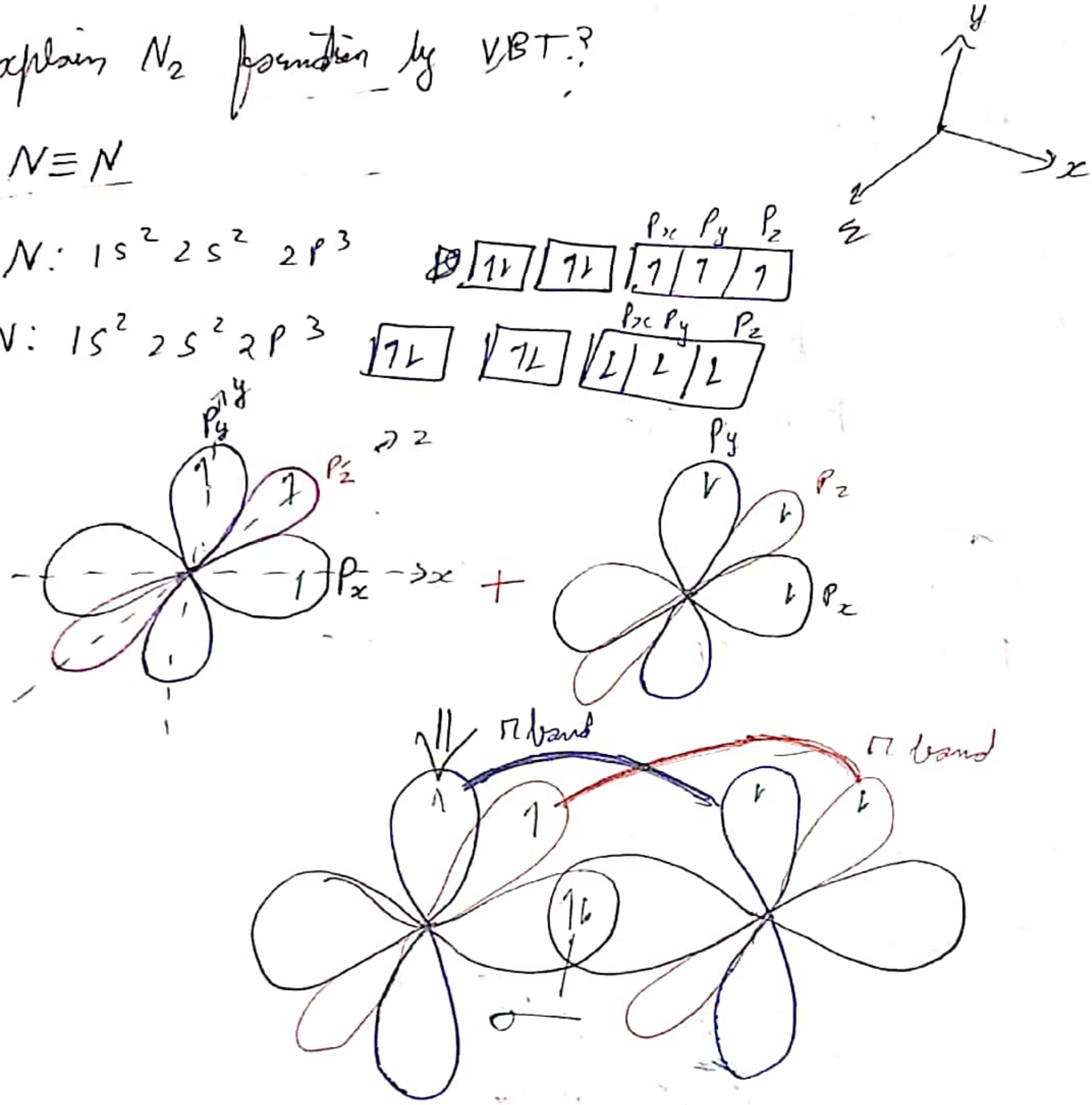


Q explain N_2 formation by VBT.?

$N \equiv N$

$N: 1s^2 2s^2 2p^3$ ⇅ ⇅ ↑ ↑ ↑ (P_x P_y P_z)

$N: 1s^2 2s^2 2p^3$ ⇅ ⇅ ↓ ↓ ↓ (P_x P_y P_z)



(2P – 2P) Overlapping

σ – bond	π – bond
1. axial overlapping	1. Co-lateral Overlapping
2. Stronger	2. Comparitively weaker
3. Can Exist alone	3. Cannot exist alone
4. Rotation around σ bond is possible	~~rotation around~~ 4. Restricted Rotation

Strength of σ – bonds :-

Case 1 when n is different (Principle quantum nos.)

$$\text{Bond Strength} \propto \frac{1}{\text{Shell no. } (n)}$$

(1s – 1s) > (2s – 2s) > (3s – 3s) σ bond strength

(2p – 2p) > (3p – 3p) > (4p – 4p)

(1s – 2p) > (1s – 3p) > (1s – 4p)

Case 2 when n is same

Bond Strength ∝ directional nature of orbital.

directional nature ⇒ s < p < d < f

n. dominates over directional nature

Q Compare bond for σ – bond

① (2s – 2s) < (2s – 2p) < (2p – 2p)

② (1s – 1s) (1s – 2s) (2s – 2p) (2p – 2p) (2s – 3s)
(3s – 3p) (3p – 3p) (3s – 4p) (4s – 4p) (3s – 4s)
(4p – 4p) (4s – 4s)

(1s – 1s) > (1s – 2s) > (2p – 2p) > (2s – 2p) > (2s – 3s) > (3p – 3p) > (3s – 3p) > (3s – 4p) > (3s – 4s) > (4p – 4p) > (4s – 4p) > (4s – 4s)

Strength of π bond

Case 1 Bond strength $\propto \dfrac{1}{\text{Shell no. } (n)}$

Case 2 Bond Strength $\propto$ directional nature.

Eg (1) $3p_\pi - 3p_\pi$ $\quad 2p_\pi - 2p_\pi \quad 2p_\pi - 3p_\pi$

$$2p_\pi - 2p_\pi > 2p_\pi - 3p_\pi > 3p_\pi - 3p_\pi$$

(2) $2p_\pi - 2p_\pi > 3p_\pi - 3p_\pi > \cancel{4p}\, 4p_\pi - 4p_\pi$

(3) $3p_\pi - 3p_\pi < 3p_\pi - 3d_\pi < 3d_\pi - 3d_\pi$

(4) $2p_\pi - 3p_\pi \quad 2p_\pi - 3d_\pi \quad 2p_\pi - 4p_\pi$

$$2p_\pi - 3d_\pi > 2p_\pi - 3p_\pi > 2p_\pi - 4p_\pi$$

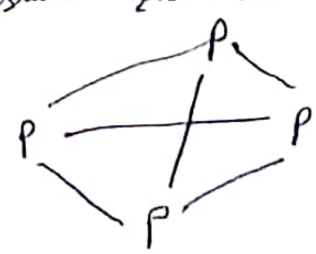
(5) $2p_\pi - 2p_\pi \quad 2p_\pi - 3d_\pi \quad 2p_\pi - 3p_\pi \quad 3p_\pi - 3p_\pi$

$$\underbrace{2p_\pi - 2p_\pi > 2p_\pi - 3d_\pi}_{\text{more stable (effective overlapping)}} > \underbrace{2p_\pi - 3p_\pi}_{\text{very less stable}} > \underbrace{3p_\pi - 3p_\pi}_{\text{Highly unstable (not effective overlapping)}}$$

Q At room temp N_2 molecule exist but P_2 does not exist?

$N \equiv N$
$2p_\pi - 2p_\pi$
effective overlapping

$P \equiv P$
$3p_\pi - 3p_\pi$
not effective so not effective
instead exist as



Q At Room ~~of~~ Temperature O_2 exist but S_2 does not why.

$O=O$

$2p_\pi - 2p_\pi$

effective overlapping

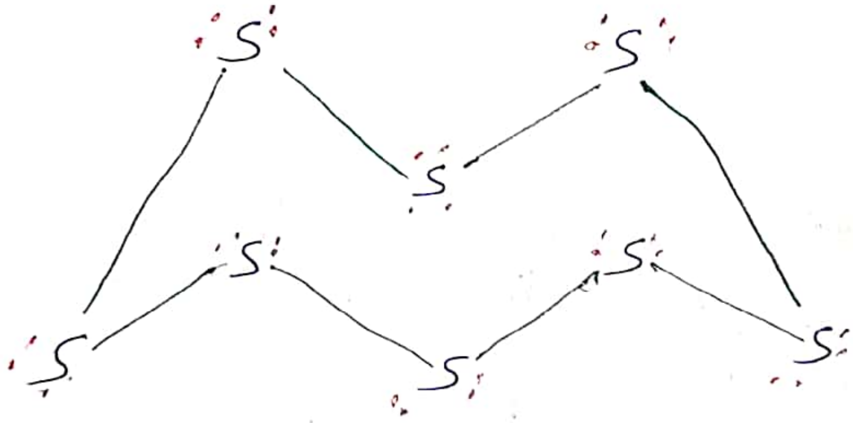
$S=S$

$3p_\pi - 3p_\pi$

not effective

So exist as S_8 (crown like structure)

[redacted]



L.P = 16

Q At room temperature CO_2 exist in monomeric form but SiO_2 molecule ~~is~~ exist in polymeric form.

$O=C=O$

$(2p_\pi - 2p_\pi)$

effective

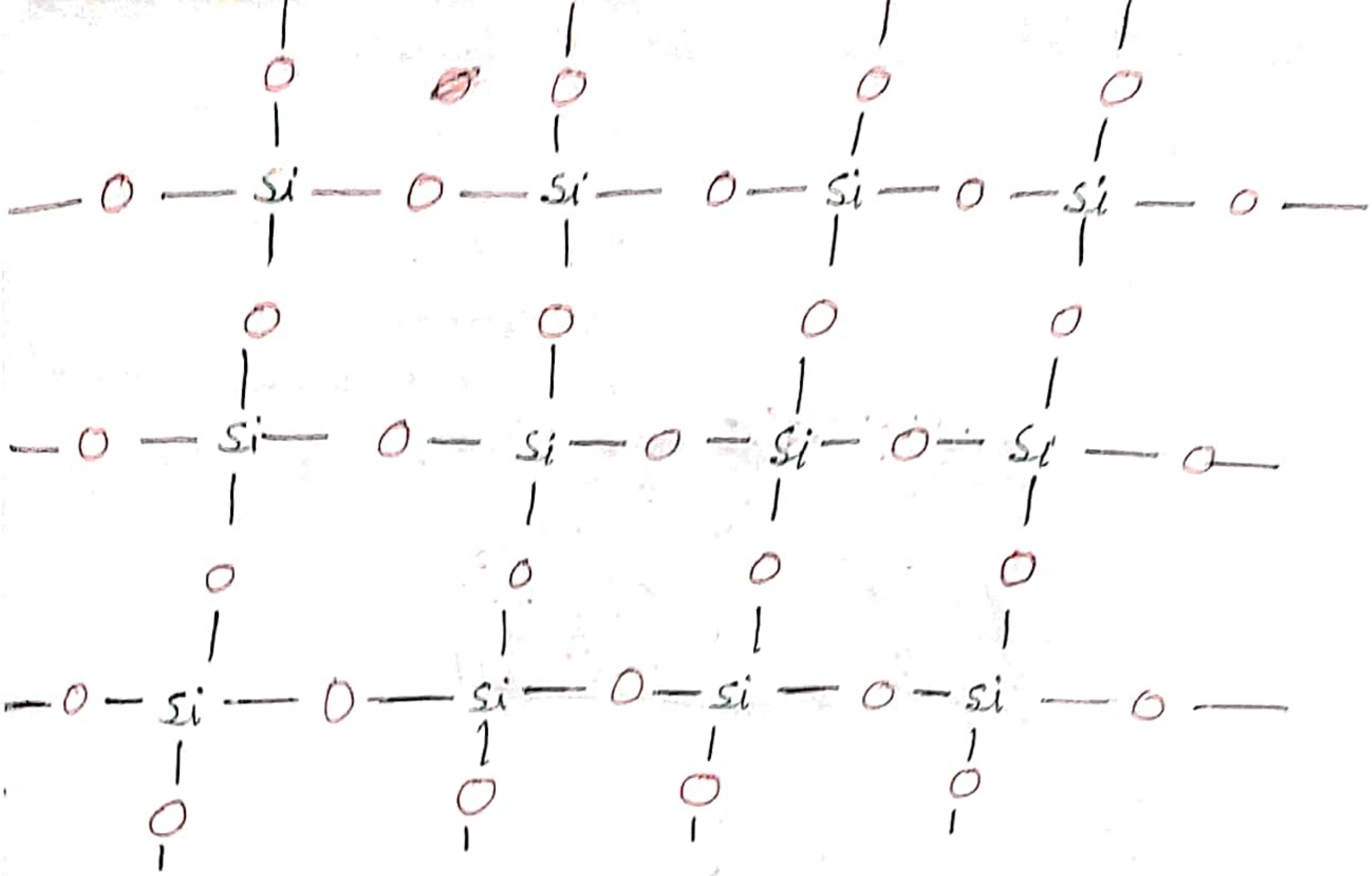
$O=Si=O$

$(2p_\pi - 3p_\pi)$

unstable

So exist in polymeric form

$(SiO_2)_n$



Bond Strength Order

Single Bond

1. $-C-C- > -Si-Si-$
2. $-N-N- < -P-P-$
3. $-O-O- < -S-S-$
4. $F-F < Cl-Cl$

due to small size L.P repulsion domenate over n

Multiple Bond

1. $-C=C- > -Si=Si-$
2. $N\equiv N > P\equiv P$
3. $O=O > S=S$

Q Bond strength order of F_2, Cl_2, Br_2 & I_2

[redacted]

expected: $F_2 > Cl_2 > Br_2 > I_2$

$\boxed{\text{Real: } Cl_2 > Br_2 > F_2 > I_2}$

Q Bond strength

H-F > H-Cl > H-Br > H-I

(1s-2p) (1s-3p) (1s-4p) (1s-5p)

Drawbacks of VBT

1. Valence Bond Theory fails to explain Geometry and ~~bond~~ bonding of molecules containing more than 2 atoms.

1. CH_4

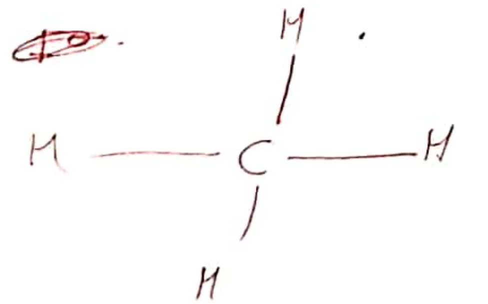
C: [↑↓] [↑|↑|]

C*: [↑] [↑|↑|↑]

H: [↑]

1 σ bond : 1s - 2s (weaker)

3σ bond : 1s - 2p (stronger)

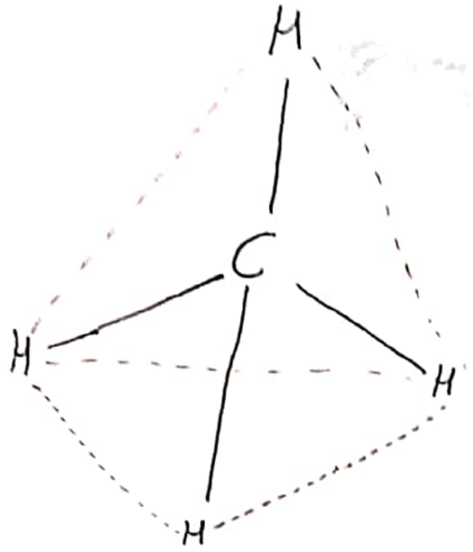


→ According to VBT CH_4 molecule has no all 4 (C-H) bond ~~bond~~ length & bond strength equal

→ But in real all 4 (C-H) bond lengths and bond strengths are same & also CH_4 has a perfect regular tetrahedral geometry.

→ To explain this, we need a new concept (hybridisation).

1. bond is weak so it should have longer bond length according to VBT but in reality all 4 bonds have a same bond length.

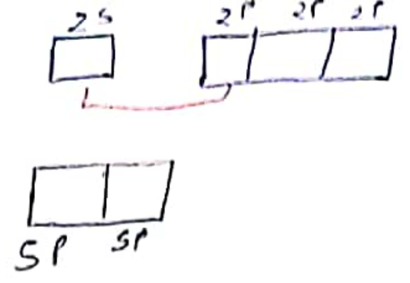


Hybridisation

→ It is the best theoritical concept to explain geometry of molecules.

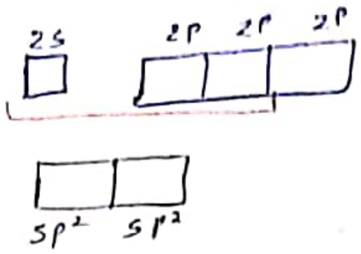
Asumptions -

1. Hybridisation is the ~~two~~ intermixing of atomic orbitals which has comparable energy or less energy difference.
2. As many no. of atomic orbitals are participated, same no. of hybrid orbitals will be formed.
3. Once the hybrid orbitals are formed they arrange themselves in such a way so that repulsion is minimised. This decides the geometry of molecule.
4. Hybrid orbitals forms σ – bond only.



Atomic Orbitals (A.O)

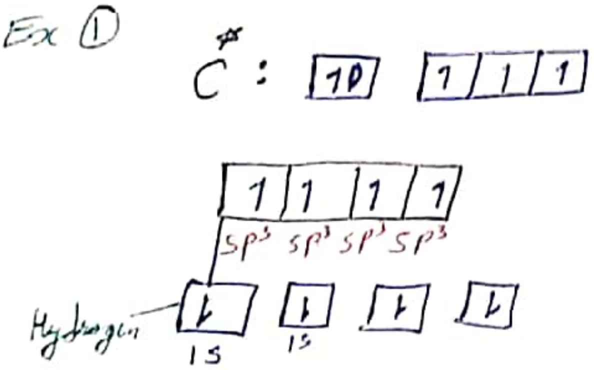
Hybrid orbital (H.O)

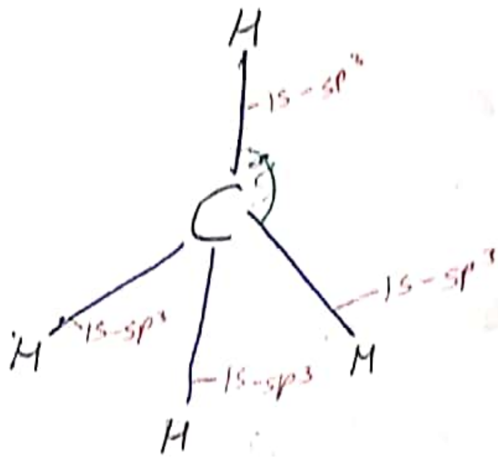


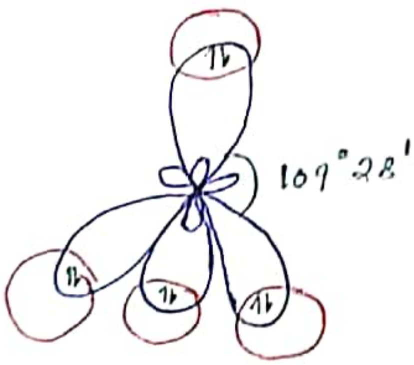
A O

H O

Ex ①



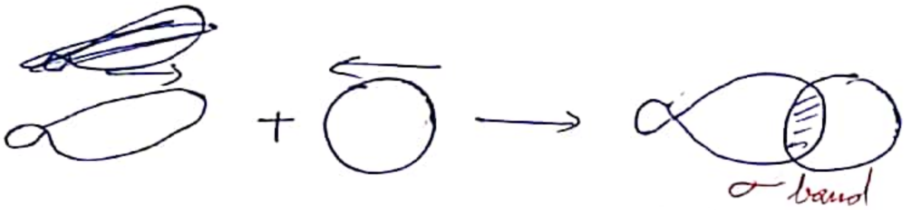


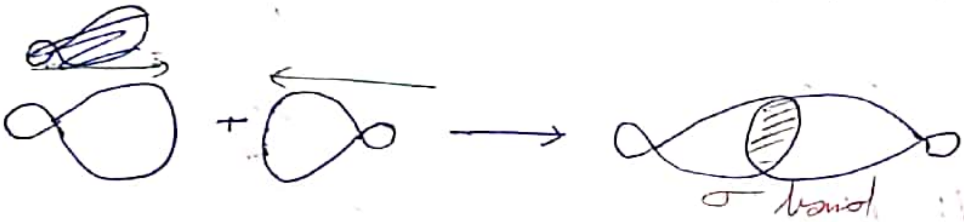


→ Here 4 (C–H) bonds are formed ⇒ This explain equal bond length, bond energy and bond strength.

→ In CH_4 all 4 bonds formed are identical with overlapping of (sp^3–1s)

Shape of Hybrid Orbital





formation of π-bond formed (not effective overlapping)

✡✡ Formed hybrid orbitals have some % s character
% p character
% d character
length & strength.

Hybridisation	% s character	% p character	Shape of H.O
SP	50%	50%	
SP^2	33.33%	66.67%	
SP^3	25%	75%	

$SP < SP^2 < SP^3$ } length of H·O $\propto$ % P character $\propto \frac{1}{\%\, S\ \text{character}}$

$SP < SP^2 < SP^3$ } Energy of H·O $\propto$ % P character $\propto \frac{1}{\%\, S\ \text{character}}$

$SP > SP^2 > SP^3$ } stability of H·O $\propto$ % S character

$\boxed{\text{Electron Negativity} \propto \%\ S\ \text{Character}}$

EN: $SP > SP^2 > SP^3$

Calculation of Hybridisation

$S + P + P + P + D + D + D$

- $S + P$: H·O = 2, SP
- $S + P + P$: H·O = 3, SP^2
- $S + P + P + P$: H·O = 4, SP^3
- $S + P + P + P + D$: H·O = 5, SP^3D
- $S + P + P + P + D + D$: H·O = 6, SP^3D^2
- $S + P + P + P + D + D + D$: H·O = 7, SP^3D^3

Steric no. (S.N.) ⇒ Total no. of Hybrid orbital

⇒ no. of σ - bond of central atom + no. of L·P on central atom.

⇒ no. of surrounding atom + no. of L.P. on central atom.

SN	Type of Hybridisation
2	sp
3	sp^2
4	sp^3
5	sp^3d
6	sp^3d^2
7	sp^3d^3

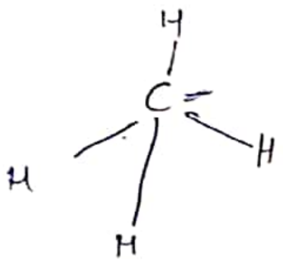
Q Find the hybridisation of central atom?

(1) $BeCl_2$

$Cl-Be-Cl$ $\quad \sigma = 2$, $L.P = 0$

$S.N. = \sigma + L.P$
$= 2+0 = 2\ (sp)$ ✓

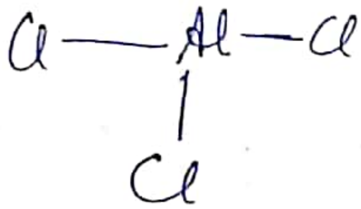
(2) CH_4



$\sigma = \cancel{3}4$
$L.P = 0$

$S.N = 4\ (sp^3)$ ✓

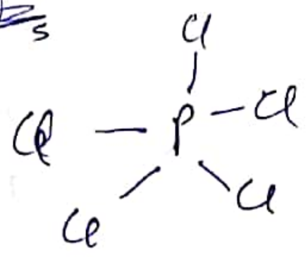
(3) $AlCl_3$



$S.N = \cancel{3+}\ 3+0$
$= 3\ (sp^2)$ ✓

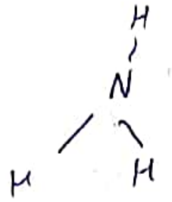
4) PCl_5
5) NH_3
6) H_2O
7) SF_6
8) IF_7
9) SCl_2
10) XeF_2
11) XeF_4

4) ~~PCl_5~~



$SN = 5\ (sp^3d)$ ✓

5)

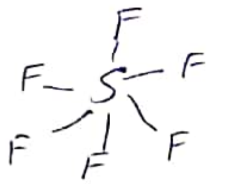


$SN = 3+1$
$= 4\ (sp^3)$ ✓

6) H — O — H

$SN = 2+2$
$= 4\ (sp^3)$ ✓

7)



$SN = 6\ (sp^3d^2)$ ✓

8)

$SN = 7\ (sp^3d^3)$ ✓

9) Cl — S — Cl

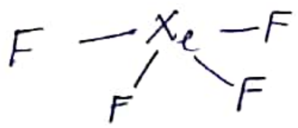
$SN = 2+2$
$= 4\ (sp^3)$ ✓

10) ~~XeF_2~~ ~~SN~~

F — Xe — F

$SN = 2+3$
$= 5\ (sp^3d)$ ✓

11) ~~XeF_4~~



$SN = 4+2$
$= 6\ (sp^3d^2)$ ✓

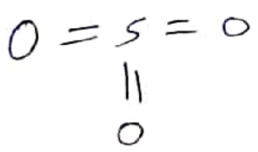
(12) SO_2
(13) SO_3
(14) XeO_4
(15) XeO_3
(16) XeO_2F_4
(17) CO_2
(18) $POCl_3$
(19) $IO_3^{\ominus}$
(20) $SnCl_3^{\ominus}$
(21) SO_3^{2-}
(22) C_2H_4
(23) C_2H_2
(24) XeO_3F_2
(25) $XeOF_4$
(26) XeF_6
(27) $XeF_5^{\oplus}$
(28) $NO_2^{\oplus}$
(29) $NO_2^{\ominus}$
(30) $XeF_5^{\ominus}$
(31) IF_7
(32) $IF_6^{\ominus}$
(33) $:I_3^{\oplus}$
(34) C_2H_6
(35) BeF_2
(36) SO_3
(37) O_3
(38) $NH_4^{\oplus}$
(39) $PH_4^{\oplus}$
(40) $ClO_4^{(-1)}$
(41) $H_3O^{\oplus}$
(42) SF_4
(43) SeF_4
(44) $TeCl_4$
(45) ClF_3
(46) $ICl_2^{(-1)}$
(47) BrF_3
(48) $BrF_2^{\ominus}$
(49) XeO_6^{4-}
(50) BrF_4^{-1}

(51) $TeCl_6$ (52) $ICl_4^{(-1)}$ (53) IF_7 (54) IF_6^- (55) XeF_6
(56) XeO_4

(12) SO_2

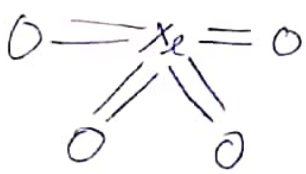
O=S=O S.N = 3 (SP²) ✓

(13) SO_3



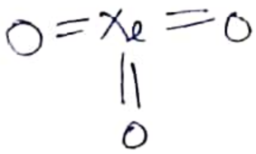
SN= 3 (SP²) ✓

(14)



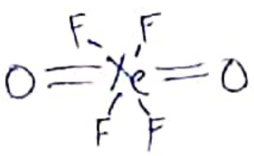
SN = 4 (SP³) ✓

(15)



SN= 4 (SP³) ✓

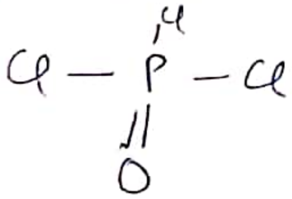
(16)

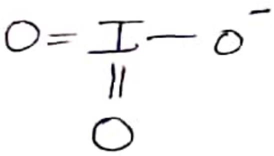


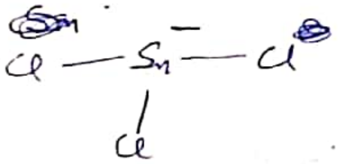
SN = 6 (SP³D²) ✓

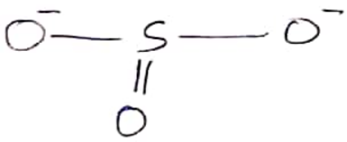
(44)

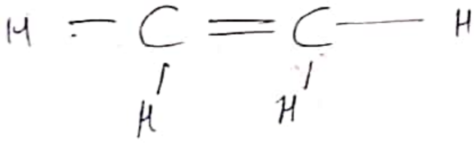
17. $O=C=O$ SN = 2 (SP) ✓

18.  SN = 4 (SP^3) ✓

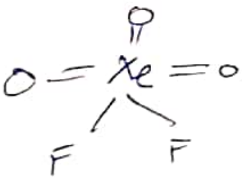
19.  SN = 4 (SP^3) ✓

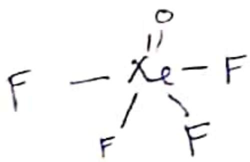
20.  SN = 4 (SP^3) ✓

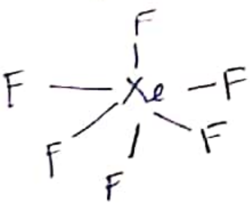
21.  SN = 4 (SP^3) ✓

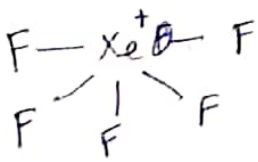
22.  SN = 3 (SP^2) ✓

23. $H-C\equiv C-H$ SN = 2 (SP) ✓

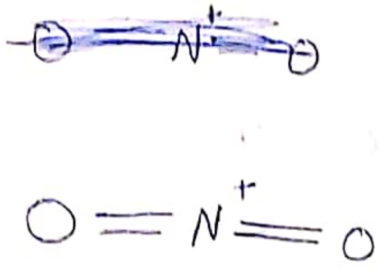
24.  SN = 5 (SP^3D) ✓

25.  SN = 6 (SP^3D^2) ✓

26.  SN = 7 (SP^3D^3) ✓

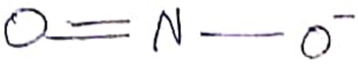
27.  SN = 6 (SP^3D^2) ✓

(28)



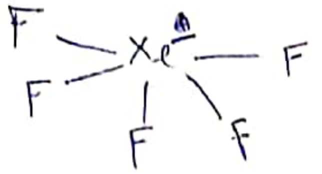
$SN = 2$ (sp)

(29)



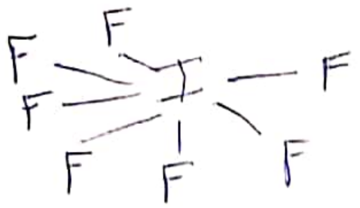
$SN = 3$ (sp^2)

(30)



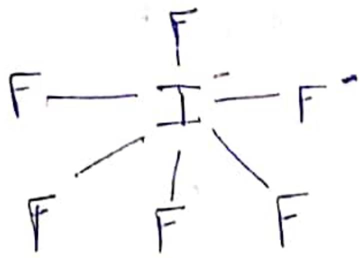
$SN = 7$ (sp^3d^3)

(31)



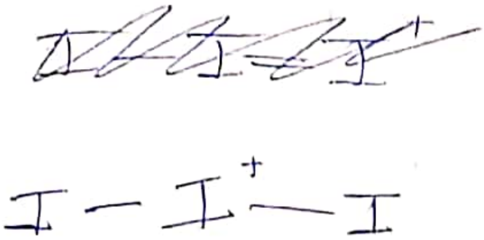
$SN = 7$ (sp^3d^3)

(32)



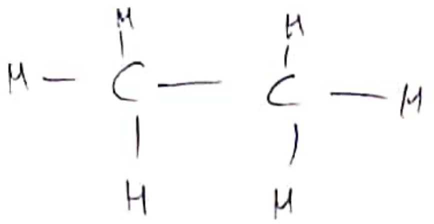
$SN = 7$ (sp^3d^3)

(33)



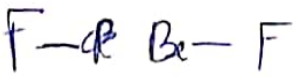
$SN = 4$ (sp^3)

(34)



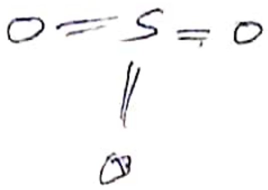
$SN = 4$ (sp^3)

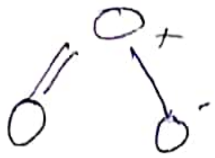
(35)

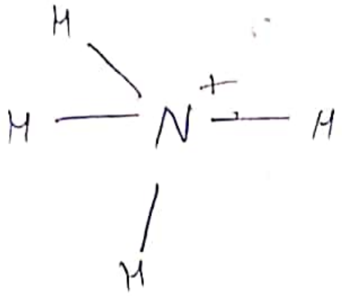


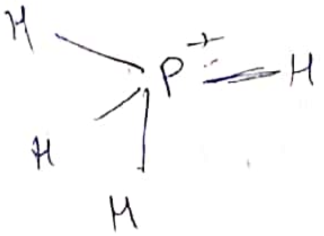
$SN = 2$ (sp)

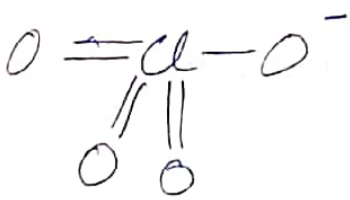
Scanned with CamScanner

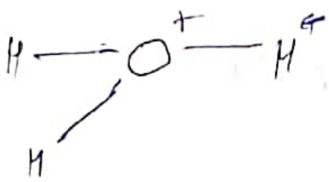
36

$SN = 3$ (sp^2) ✓

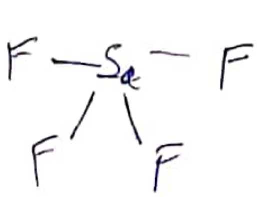
37

$SN = 3$ (sp^2) ✓

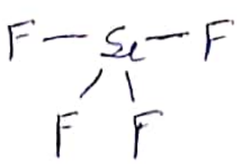
38

$SN = 4$ (sp^3) ✓

39

$SN = 4$ (sp^3) ✓

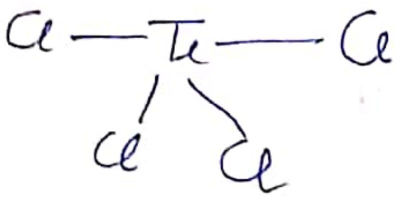
40

$SN = \cancel{7}\,4$ (sp^3) ✓

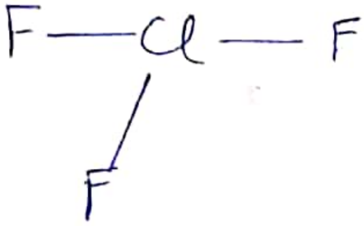
41

$SN = 4$ (sp^3) ✓

42

$SN = 5$ (sp^3d) ✓

43

$SN = 5$ (sp^3d) ✓

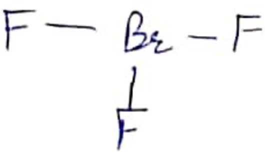
44 

44  SN = 5 (sp^3D) ✓

45  SN = 5 (sp^3D) ✓

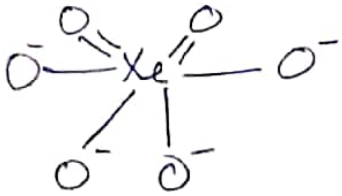
46 ~~Cl — I — Cl~~

46 $Cl - I^- - Cl$ SN = 5 (sp^3D) ✓

47  SN = 5 ~~(sp^2)~~ (sp^3D) ✓

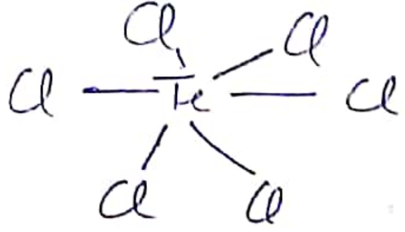
48 $F - Br^- - F$ SN = 5 ~~(sp^2)~~ (sp^3D) ✓

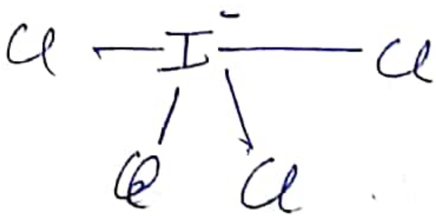
49 ~~XeO_6^{4-}~~

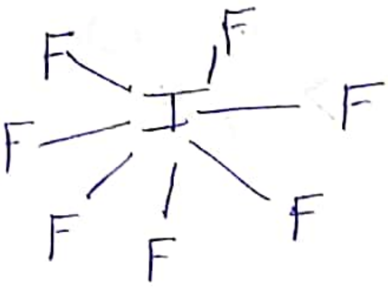
 SN = 6 (sp^3D^2) ✓

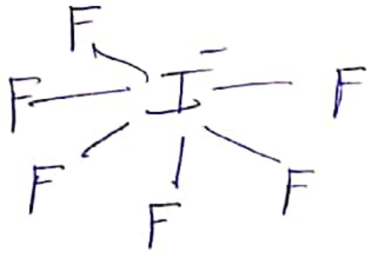
50  SN = 6 (sp^3d^2) ~~(sp^3)~~ ✓

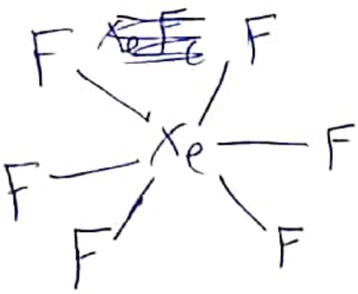
48

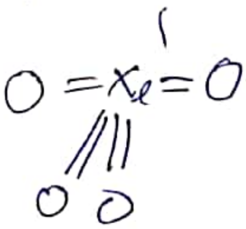
51  $SN = 6$ (sp^3d^2)

52  $SN = 6$ (sp^3d^2)

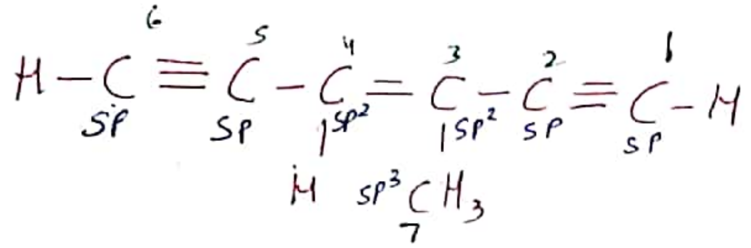
53  $SN = 7$ (sp^3d^3)

54  $SN = 7$ (sp^3d^3)

55  $SN = 7$ (sp^3d^3)

56  $SN = 4$ (sp^3)

Q EN order ?



EN ∝ % s character

EN: $sp > sp^2 > sp^3$

$(C_1 = C_2 = C_5 = C_6) > (C_4 = C_3) > C_7$ [EN order]

Q 2. Compare C–C bond strength

C_2H_6 C_2H_4 C_2H_2

sp^3 sp^2 sp

$C_2H_6 < C_2H_4 < C_2H_2$

Q 3. Compare C–C bond length

C_2H_6 C_2H_4 C_2H_2

sp^3 sp^2 sp

$C_2H_6 > C_2H_4 > C_2H_2$

Q 4. Compare EN of the central atom.

C_2H_6 C_2H_4 C_2H_2

sp^3 sp^2 sp

$C_2H_6 < C_2H_4 < C_2H_2$

Structure of Different Hybridisation (for no L.P on central atom).

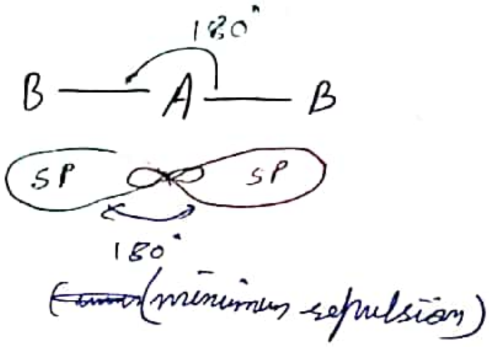
① SP-Hybridisation -

S + (any P orbital)

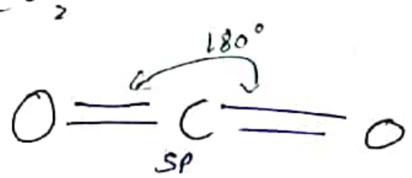
S + (P_x/P_y/P_z)

Geometry : linear (planar)

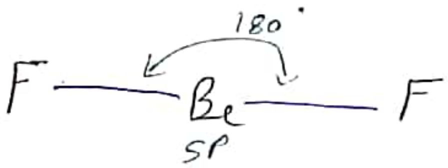
Angle : 180°



Eg CO_2



Eg ② BeF_2

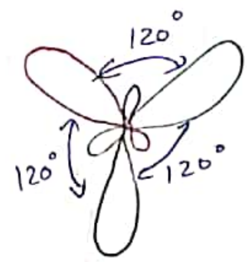


② SP^2- Hybridisation -

S + (any 2 P Orbitals)

~~S + (2 P_x/2P_y/2P_z~~

S + P + P

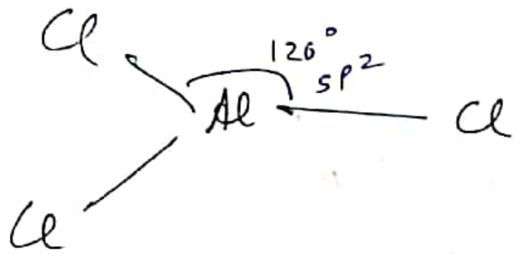


Geometry : Trigonal Planar

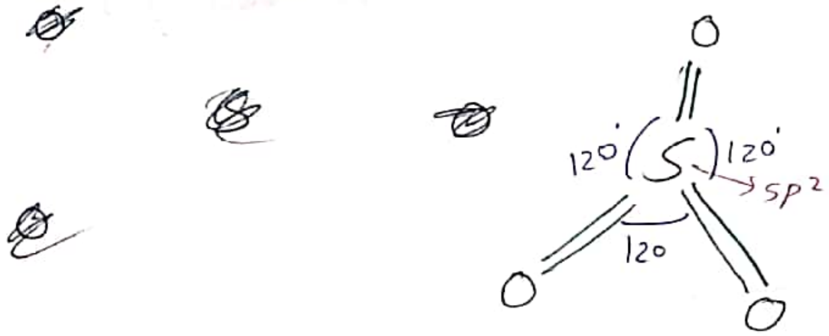
Bond Angle : 120° (3)

(Planar)

Eg ① $AlCl_3$



② SO_3

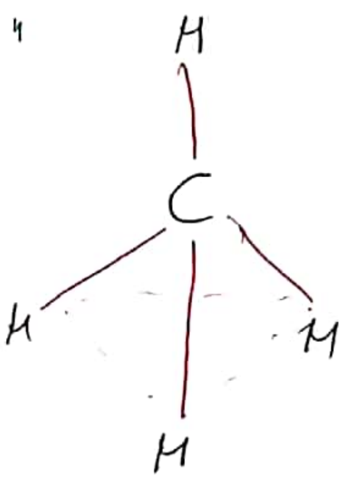


③ SP^3 – Hybridisation –

$(s + P_x + P_y + P_z)$

Geometry : Tetrahedral
Bond Angle : 109.5°/109°28'
(non-Co-Planar)

Eg ①. CH_4



6 angles of 109.5°

(4) sp^3d Hybridisation

$s + p_x + p_y + p_z + d_{z^2}$

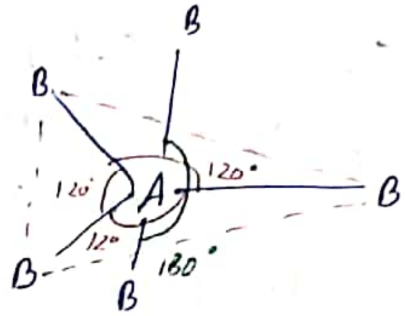
Geometry – Trigonal Bipyramidal

Bond Angle – 120° (3)

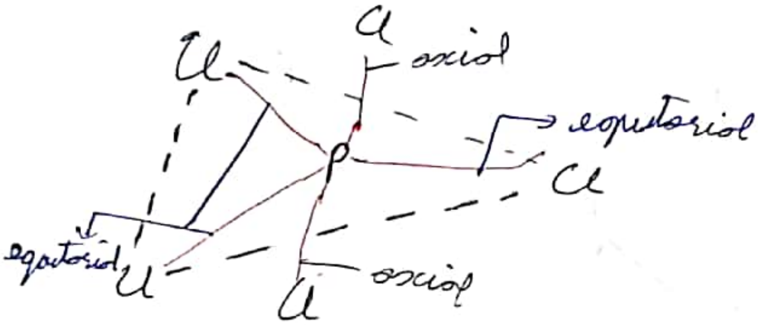
90° (6)

180° (1)

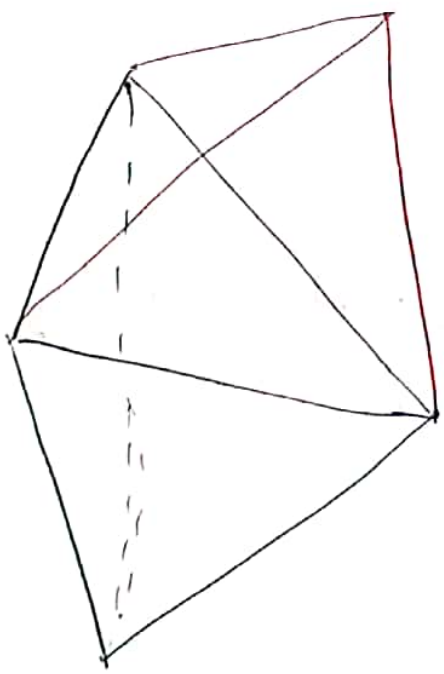
Non-Planar



Eg ① PCl_5



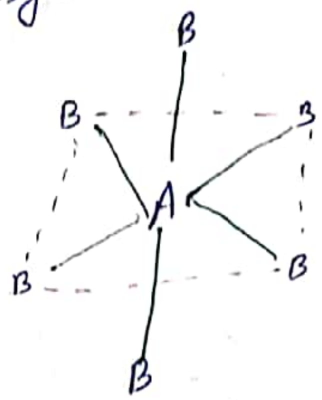
~~120~~



5. sp^3d^2 Hybridisation

$(s + p_x + p_y + p_z + d_{z^2} + d_{x^2-y^2})$

Geometry :- Octahedral / Square Bi Pyramidal

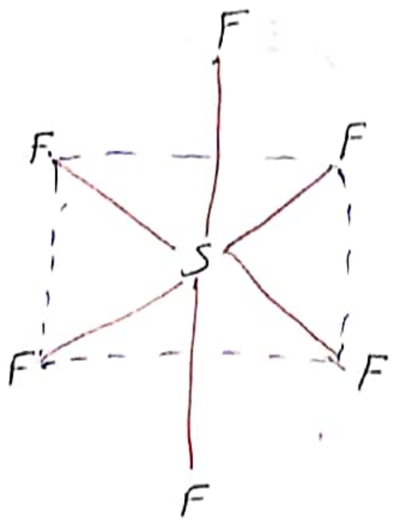


Band Angle :- 90° (12)

180° (3)

Non-Planar

Eg 1. SF_6



6. sp^3d^3 Hybridisation

$(s + p_x + p_y + p_z + d_{z^2} + d_{x^2-y^2} + d_{xy})$

Geometry :- Pentagonal Bi Pyramidal

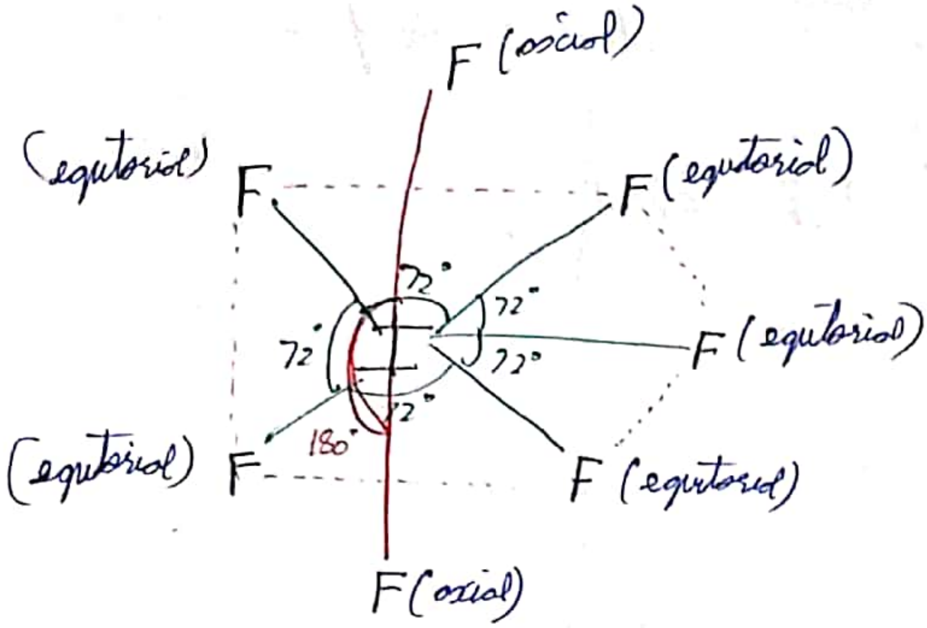
Bond Angle :- 180° (1)

90° (10)

72° (5)

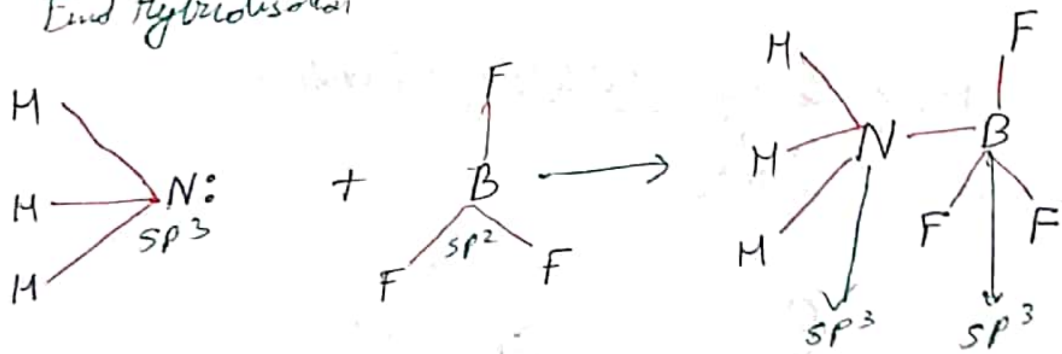
144° (5)

Non-Planar

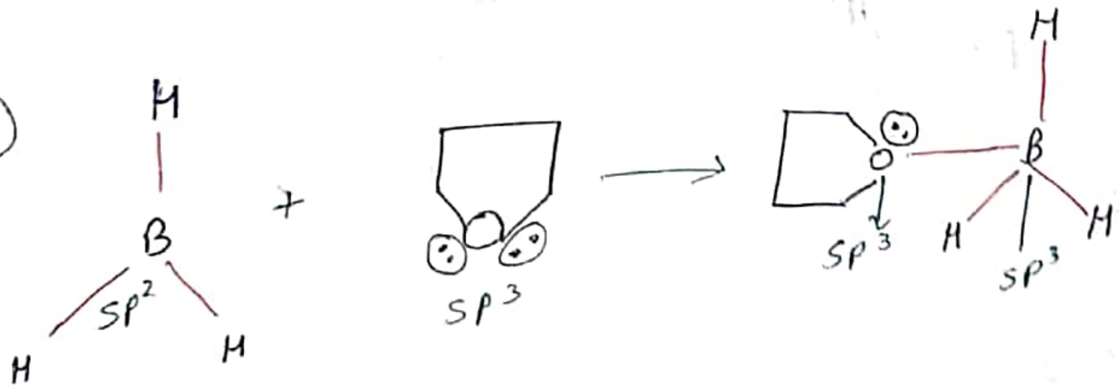


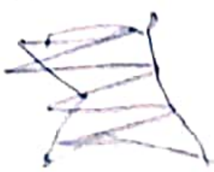
Find Hybridisation

(1)



(2)





Q In sp^3d hybridisation which of the following orbitals may be involved.

a) d_{xy} ✓b) d_{z^2} c) $d_{x^2-y^2}$ D) both (B) & (C)

Q Select the correct Statement.

✓(1) sp^2

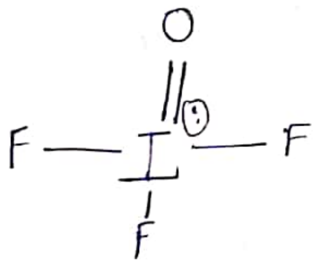
✓(2) sp^3

(3) ns + any 2 - nd + 3 - np → sp^3d^2

(4) ns + 3-np + 2(n-1)d → sp^3d^2

✓(5) ns + 3 np + 2 axial (n-1) d → sp^3d^2

Q which of the following orbitals are involved in hybridisation of central atom of IOF_3.



SN = 1 + 4
= 5

✓a) s

✓b) P_z

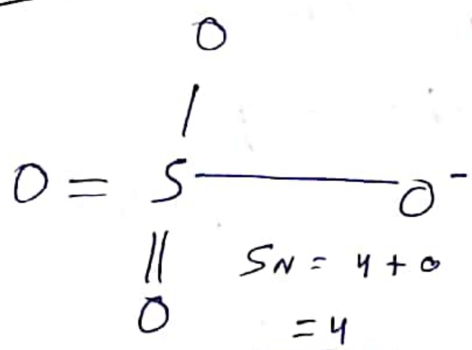
c) $d_{x^2-y^2}$

✓d) d_{z^2}

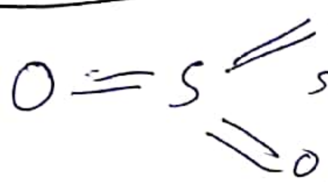
Q. Match the column:

species	Atomic orbital involved in hybridisation.
A) SO_4^{2-}	(P) s
B) SO_3 (xz-plane)	(Q) P_z
c) XeO_2F_2	(R) d_{xy}
D) XeF_4	(S) $d_{x^2-y^2}$
e) SF_2Cl_2	(T) d_{z^2}

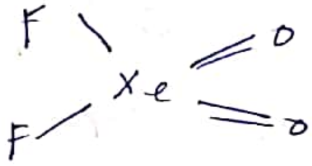
A – PQ
B – PQ
C – PQT
D – PQTS
E – PQT



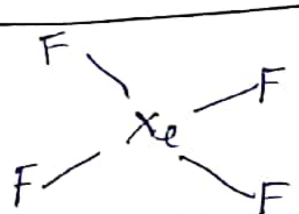
$SN = 4 + 0$
$= 4$
$\boxed{sp^3}$ – A (s, p_x, p_y, p_z)



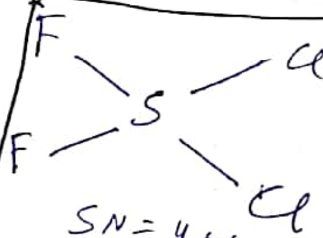
$SN = 3$
$\boxed{sp^2}$ – B
(s, p_x, p_z) (xz plane)



$SN = 1 + 4$
$= 5$
$\boxed{sp^3d}$ – C $(s, p_x, p_y, p_z, d_{z^2})$



$SN = 4 + 2$
$= 6$
$\boxed{sp^3d^2}$ – D
$(s, p_x, p_y, p_z, d_{z^2}, d_{x^2-y^2})$



$SN = 4 + 1$
$= 5$
$\boxed{sp^3d}$ – E
$(s, p_x, p_y, p_z, d_{z^2})$

VSEPR THEORY

(Valence Shell e^- pair repulsion theory)

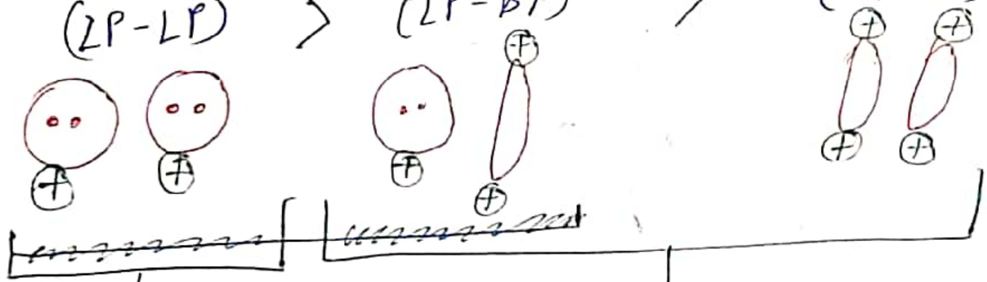
→ According to VSEPR theory, Bond Pair & Lone Pairs repel each other, To minimize these repulsions, molecule occures such shape/structure in which repulsion between e^- pairs are minimum.

→ Minimum repulsion results in greater stability.

→ Repulsion is of 3 types.

1. LP-LP
2. LP-BP
3. BP-BP

(LP-LP) > (LP-BP) > (BP-BP)



→ Here the Size of Lone Pair ~~e~~ e^- cloud is larger due to absence of 2^{nd} nucleus.

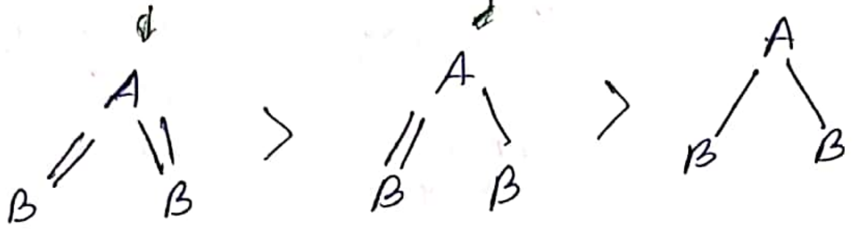
→ Hence the distance btw two lone pair is less & repulsion is high

→ Bond Pair get ~~fl~~ flattened due to streaching by two nucleus. Hence the distance in (LP-BP) or (BP-BP) is increased & repulsion is lower.

Types of Bond Pair repulsion

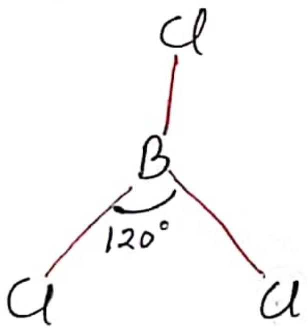
1. Multiple Bond - Multiple Bond
2. Multiple Bond - Single Bond
3. Single Bond - Single Bond

(MB-MB) > (MB-SB) > (SB-SB)

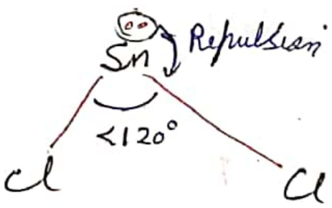


→ In Multiple bond, π electrons occupy more space around central atom so their repulsion is higher as compared to single bond

Eg ① sp^2

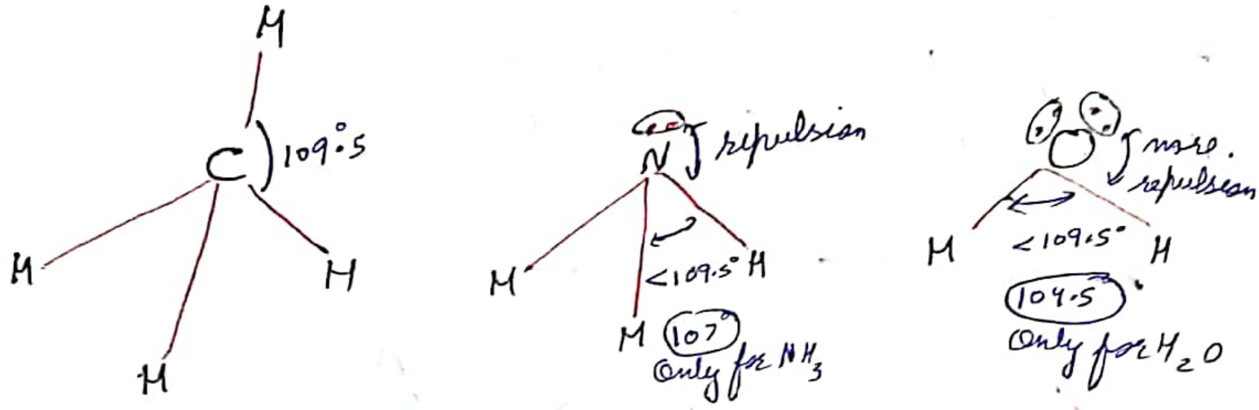


$\sigma = 3$
l.P = 0
SN = 3



$\sigma = 2$
l.P = 1
SN = 3

② SP^3



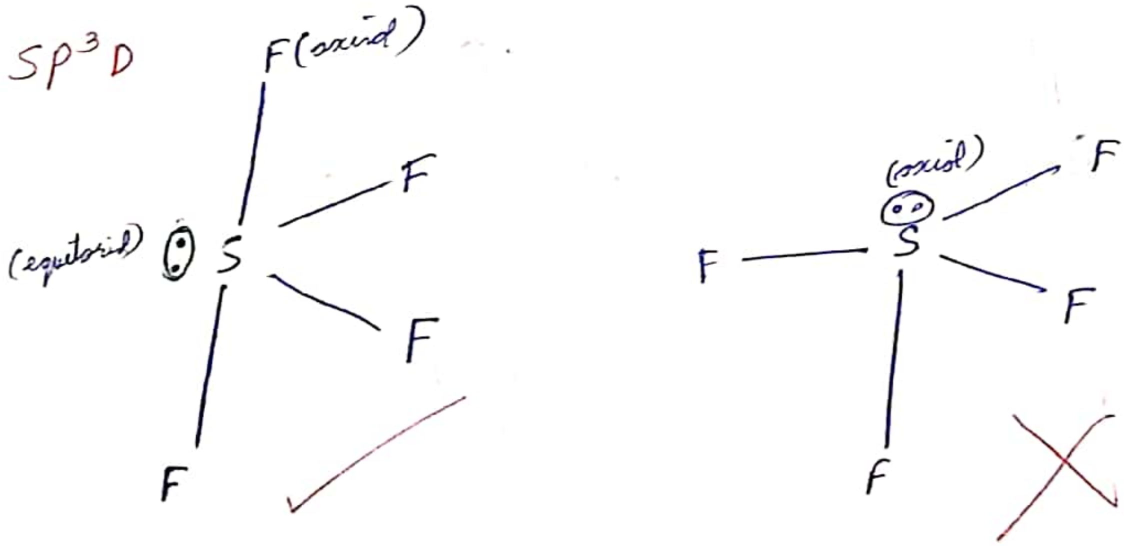
$\sigma = 4$
$L.P = 0$
$SN = 4$

$\sigma = 3$
$L.P = 1$
$SN = 4$

$\sigma = 2$
$L.P = 2$
$SN = 4$

→ Bond angle of Ammonia & H_2O should be 109.5° according to hybridisation but due to L.P–B.P repulsion, In ammonia, the angle decreases. Similarly H_2O has 2 lone Pairs (L.P) & hence more repulsion so decrease in Bond Angle will be more.

③ SP^3D

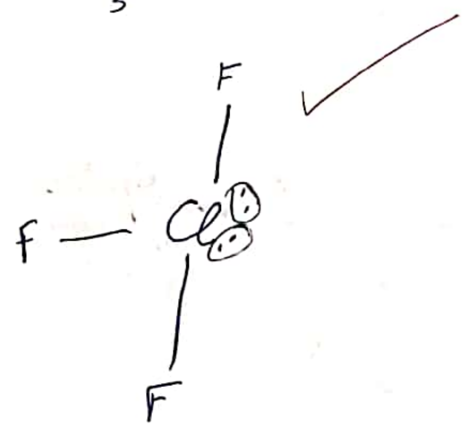


SF_4 $(\sigma = 4, L.P = 1)$

at 90°
L.P – B.P = 2
& less repulsion

at 90°
L.P – B P = 3

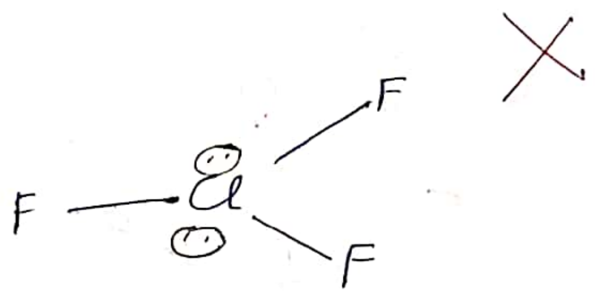
ClF_3 $(\sigma = 3, LP = 2)$



at 90°

L.P - BP = 4

less Repulsion

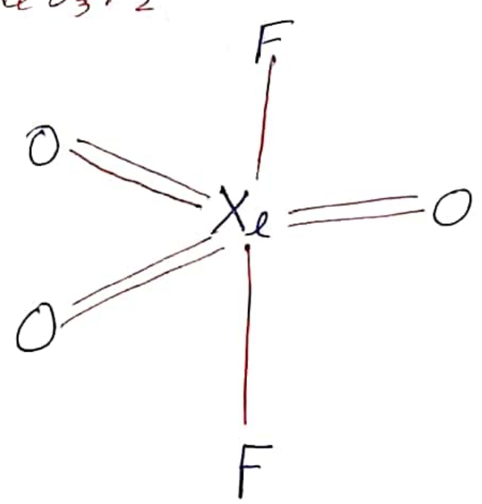


at 90°

LP - BP = 6

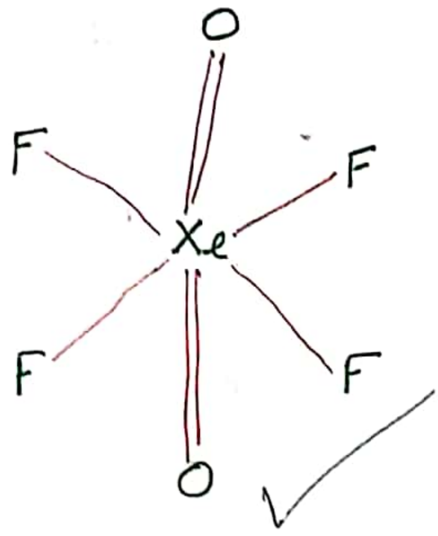
⇒ Lone Pair & Double bond will always be at equitorial location.

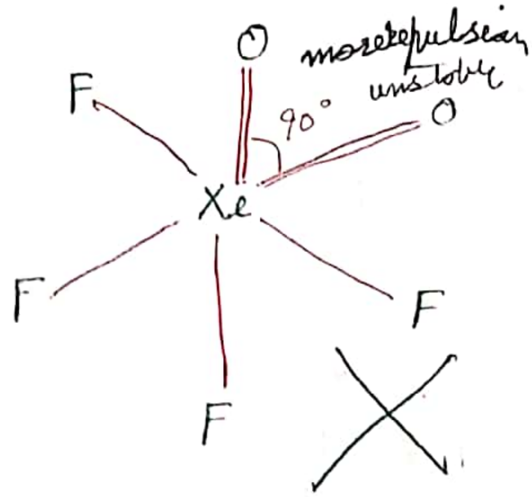
~~XeO~~ XeO_3F_2



(4) sp^3d^2 (Put Lone Pair & Double bond at 180°)

XeO_2F_4



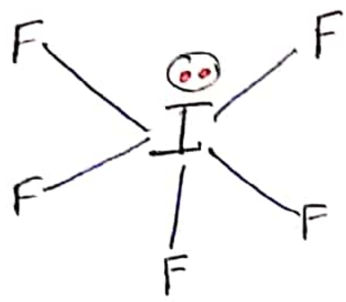


$\sigma = 6$

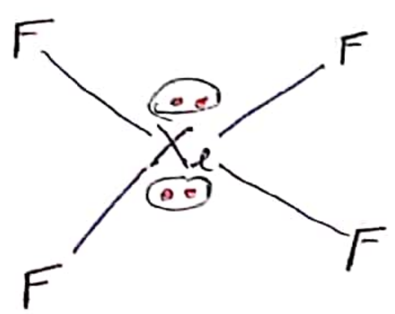
sp^3d^2

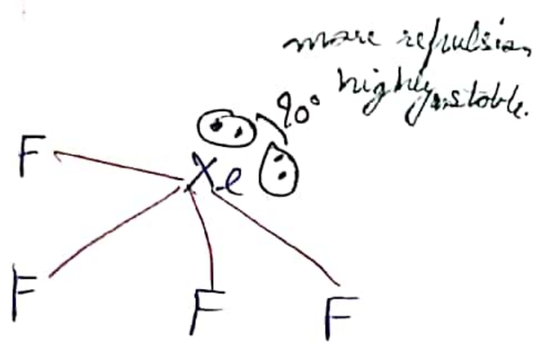
Square bipyramidal

Eg (1) IF_5



(2) XeF_4





Q Calculate

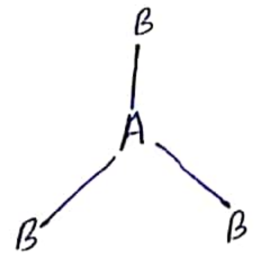
① σ - bond ② Lone Pair ③ No. of hybrid orbitals (H·O)
④ Hybridisation ⑤ e^- geometry ⑥ shape/geometry / structure
⑦ Bond Angle ⑧ Planar / non-Planar ⑨ Max no. of atoms in one plane

SP
① AB_2 (~~SP~~) eg. $BeCl_2$

B — A — B

① σ bond = 2
② Lone Pair = 0
③ No. of H·O· = 2
④ Hybridisation = SP
⑤ e^- geometry = linear
⑥ Shape = linear
⑦ Bond angle = 180° (1)
⑧ Planar
⑨ max no. of atoms in a plane = 3

SP^2
② AB_3 eg. $AlCl_3$

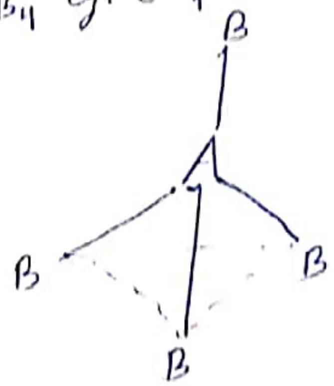


① σ = 3
② L·P = 0
③ H·O = 3
④ SP^2
⑤ Trigonal Planar
⑥ Trigonal Planar
⑦ Bond Angle — 120° (3)
⑧ Planar
⑨ 4

(3) AB_2E_{1-2} represent a lone Pair
eg. $SnCl_2$

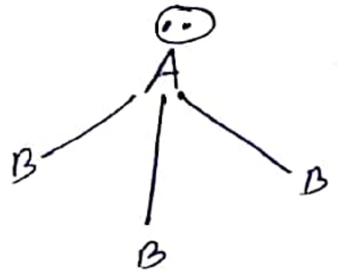
1. $\sigma = 2$
2. $L.P = 1$
3. $H.O = 3$
4. sp^2
5. Trigonal Planar
6. V-Shape / Bent Shape / Angular
7. $<120°$ (1)
8. Planar
9. 3

(4) sp^3 AB_4 eg. CH_4



1. $\sigma = 4$
2. $L.P = 0$
3. $H.O = 4$
4. sp^3
5. Tetrahedral
6. Tetrahedral
7. 109.5° (6)
8. Non-Planar
9. 3

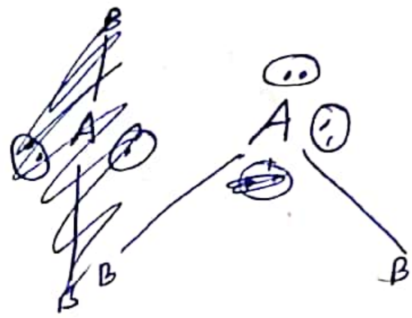
5 AB_3E_1 eg NH_3



1. $\sigma = 3$
2. L.P = 1
3. H.O = 4
4. sp^3
5. Tetrahedral
6. Trigonal Pyramidal
7. $< 109.5°$ (3)
8. Non - Planar
9. 3

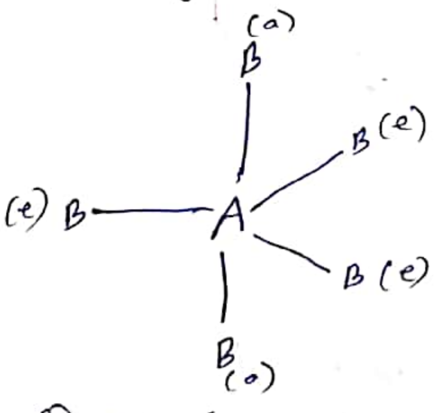
sp^3d

6 AB_2E_2 H_2O



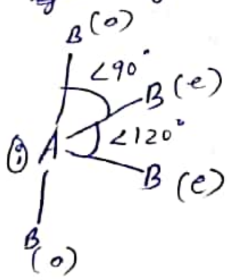
1. $\sigma = 2$
2. L.P = 2
3. H.O = 4
4. sp^3
5. Tetrahedral
6. ~~Linear~~ V-Shape
7. $< 109.5°$ (1)
8. Planar
9. 3

7 AB_5 eg PCl_5



1. $\sigma = 5$
2. LP = 0
3. H.O = 5
4. sp^3d
5. Trigonal Bipyramidal
6. Trigonal Bipyramidal
7. 120° (3)
   90° (6)
   180° (1)
8. Non - Planar
9. 4

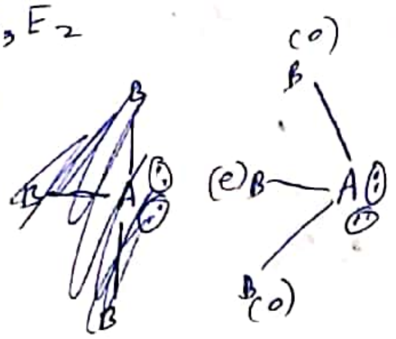
8 AB_4E_1 eg. SF_4



1. $\sigma = 4$
2. L.P = 1
3. H.O = 5
4. sp^3d
5. Trigonal Bipyramidal
6. See - saw
7. $< 90°$ (4)
   < 180 (1)
   $< 120°$ (1)
8. Non - Planar
9. 3

(9) AB_3E_2

eg. ClF_3



(1) $\sigma = 3$

(2) $L \cdot P = 2$

(3) $H \cdot O = 5$

(4) sp^3d

(5) Trigonal Bipyramidal

(6) Distorted T-Shape

(7) $< 90^\circ (2)$ $\leq 180^\circ (1)$ a-a

(8) ~~Non~~-Planar

(9) 4

sp^3d^2

(10) AB_2E_3 eg XeF_2

(1) $\sigma = 2$

(2) $L \cdot P = 3$

(3) $H \cdot O = 5$

(4) sp^3d

(5) Trigonal Bipyramidal

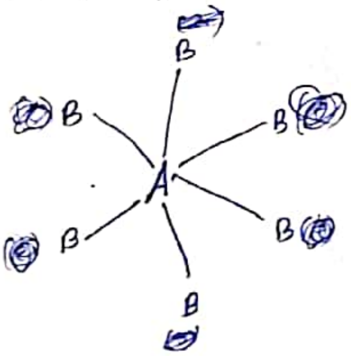
(6) Linear

(7) $180^\circ (1)$ a-a

(8) Planar

(9) 3

(11) AB_6 eg SF_6



(1) $\sigma = 6$

(2) $L \cdot P = 0$

(3) $H \cdot O = 6$

(4) sp^3d^2

(5) Square Bi Pyramidal

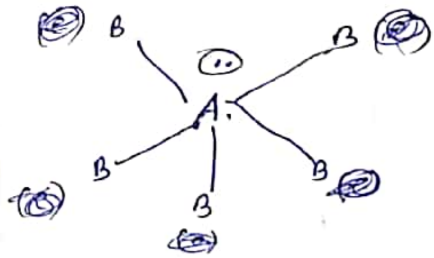
(6) Square Bi Pyramidal

(7) $90^\circ (12)$
$180^\circ (3)$

(8) Non-Planar

(9) 5

(12) AB_5E_1 eg. BrF_5



(1) $\sigma = 5$

(2) $L \cdot P = 1$

(3) $H \cdot O = 6$

(4) sp^3d^2

(5) Square Bi Pyramidal

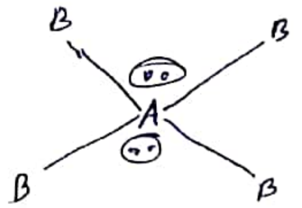
(6) Square Pyramidal

(7) $< 90^\circ (8)$ – around 70°
$< 180^\circ (2)$ – around 175°

(8) Non-Planar

(9) 4

(13) AB_4E_2 eg XeF_4

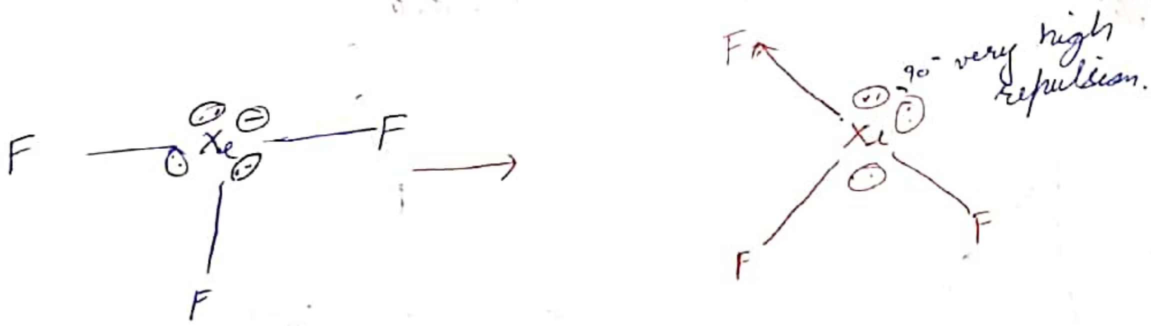


(1) $\sigma = 4$
(2) $L \cdot P = 2$
(3) $H \cdot O = 6$
(4) sp^3d^2
(5) Square Bipyramidal
(6) ~~Sq~~ Square Planar
(7) $90°(4)$ $180°(2)$
(8) Planar
(9) 5

(14) AB_3E_3

at 90°, L·P-LP repulsion is so much that it is unstable and so geometry does not exist.

Q $XeF_3^{\ominus}$ does not exist why



$L \cdot P = 3$
$\sigma = 3$
$H \cdot O = 6$
sp^3d^2
L·P-LP repulsion is so high that it is unstable

H.W. 04-06-2024

O-1 (8-17)

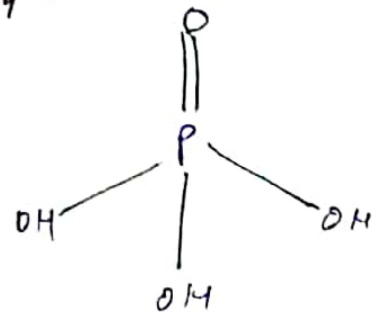
O-2 (6-11)

Roce - 13, 14

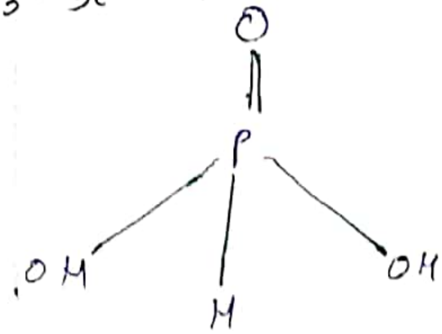
H.W. 06-06-2024

JM - 1, 2, 4, 5, 7, 8, 14, 16, 23, 24, 26, 27, 29, 31, 33, 42, 45, 46, 47, 53, 56, 57, 58

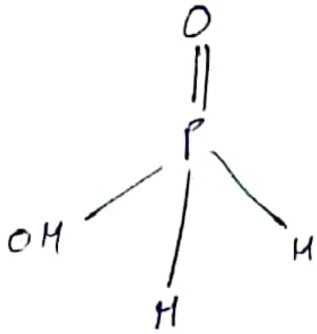
H_3PO_4 (Phosphoric acid)



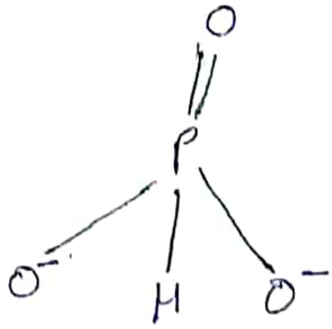
H_3PO_3 (Phosphorus acid)



H_3PO_2 (Hypophosphorus acid)



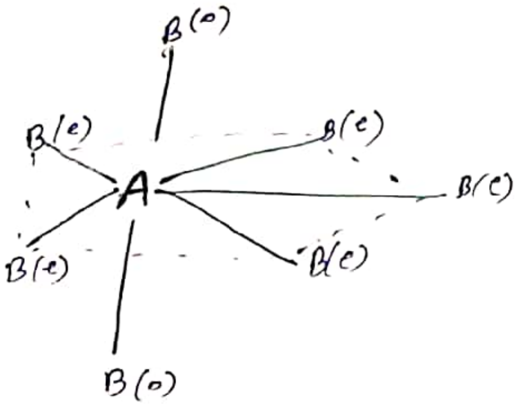
$H_3PO_4 \xrightarrow[-H_2O]{} HPO_3^{2-}$



Note – Whenever L·P–L·P repulsion is present at 90° in SP^3D^2 hybridisation, structure does not exist.

SP^3D^3

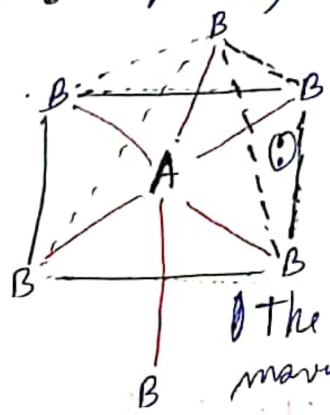
(15) AB_7 eg IF_7



1. $\sigma = 7$
2. $L \cdot P = 0$
3. $H \cdot O = 7$
4. SP^3D^3
5. Pentagonal Bi Pyramidal
6. Pentagonal Bi Bipyramidal
7. 72°(5) – (e – e)
   90°(10) – (a – e)
   180°(1) – (a – a)
   144°(5) – (e-e)
8. Non-Planar
9. 6

Yaad Rakhna Hai

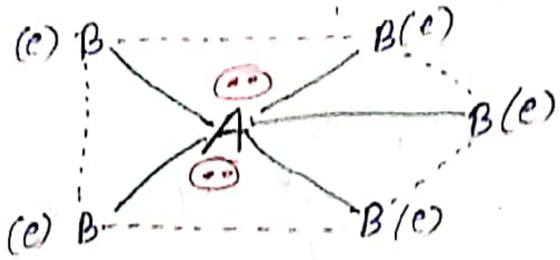
(16) AB_6E eg. XeF_6, IF_6^-, $XeOF_5^-$



The lone pair keeps moving in the Block From Pyramid Side

1. $\sigma = 6$
2. $L \cdot P = 1$
3. $H \cdot O = 7$
4. SP^3D^3
5. Pentagonal BiPyramidal
6. Distorted octahedral
   or
   Capped octahedron
7. ×
8. Non-Planar
9. ×

(17) AB_5E_2 eg. XeF_5^-



(1) $\sigma = 5$
(2) L.P = 2
(3) H.O = 7
(4) sp^3d^3
(5) Pentagonal Bipyramidal
(6) ~~Planar~~ Pentagonal Planar

sp^3d^3 में L.P हमेशा
180° पर होगा।

(7) 72° (5) - (e-e)
144° (5) - (e-e)

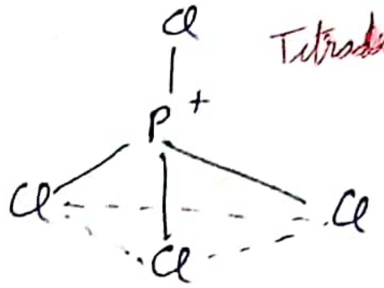
(8) Planar

(9) 6

Q ~~De~~ Draw Structure

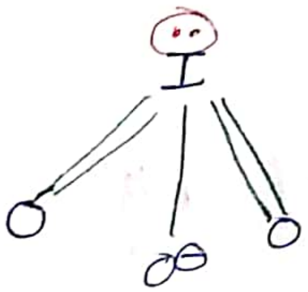
(1) PCl_4^+
(2) IO_3^-
(3) SO_2
(4) XeO_2F_2
(5) ICl_4^-
(6) $SOCl_2$
(7) Br_2F_5

① PCl_4^+ (sp^3)



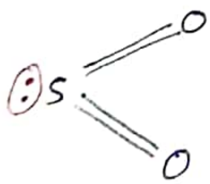
Tetrahedral

② IO_3^- (sp^3)



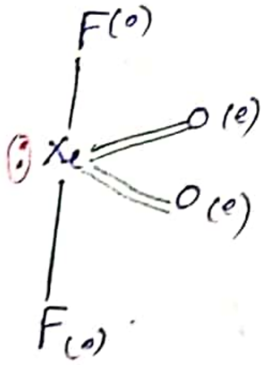
Pyramidal

③ SO_2 (sp^2)



V-shape

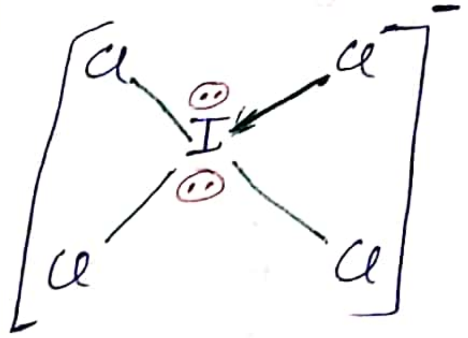
④ XeO_2F_2 (sp^3d)



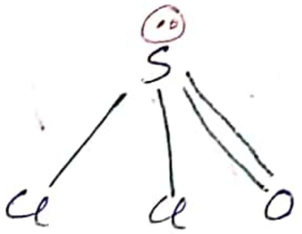
See-saw

⑤ ICl_4^- (sp^3d^2)

Square Planar

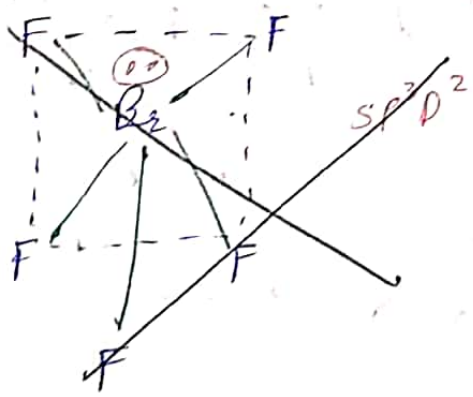


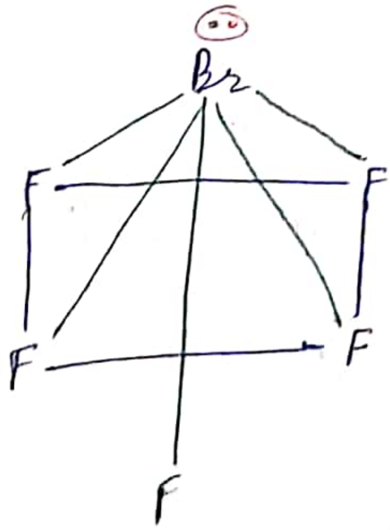
⑥ $SOCl_2$ (sp^3)



Pyramidal

⑦ BrF_5 (sp^3d^2)



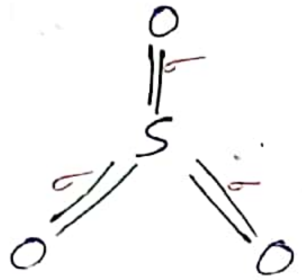


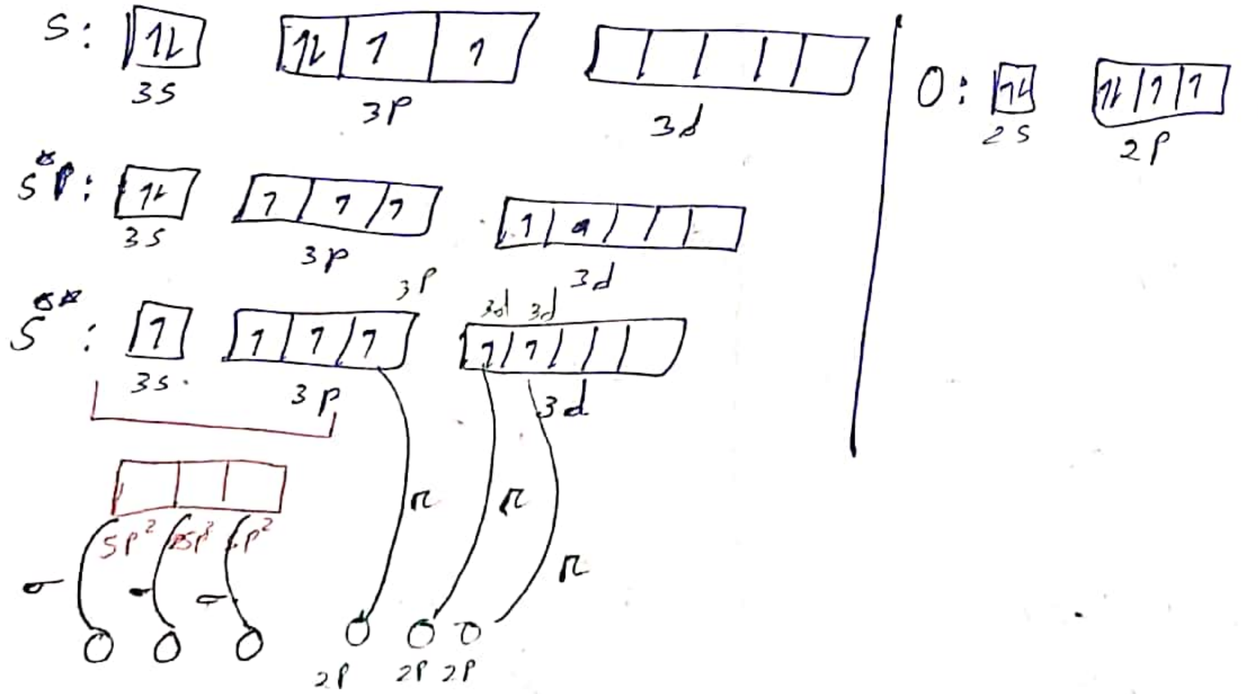
Square Pyramidal

Q. SO_3 How many π-bond & Types of π-bonds

Type of π-bond Calculation

sp^2



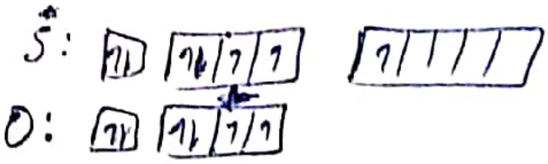


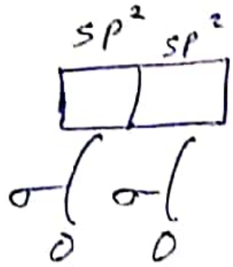
$\boxed{\begin{array}{l} p_\pi - p_\pi \to 1 \text{ Bond} \\ p_\pi - d_\pi \to 2 \text{ Bond} \end{array}}$

Note :- Oxygen atom always ~~form~~ use 2p orbital to for π-bond

Q1. SO_2 how many π-bonds & types of π-bonds.

(1) SO_2





$p\pi - p\pi \rightarrow$ 1 Bond ✓

$p\pi - p\pi \rightarrow$ 1 Bond ✓

(2) ClO_2^{+}

(3) XeO_4

(4) ClO_4^{-}

(5) C_2H_2

(6) CO_3^{2-}

(7) ClO_3^{-}

Q2. No. of species that have more $p\pi - p\pi$ than $p\pi - p\pi$ bonds :-

SO_2, NO_3^-, SO_3, XeO_2F_4, PO_4^{3-}, IO_2^-

$COCl_2$, NO_2^-, CO, NOF_3

Q1 ② ClO_2^{+} (sp^2)

Cl°: [↑↓] [↑ | ↑ | ↑] [↑ | ↑ | | |]

O: [↑↓] [↑↓ | ↑ | ↑]

[· | |]

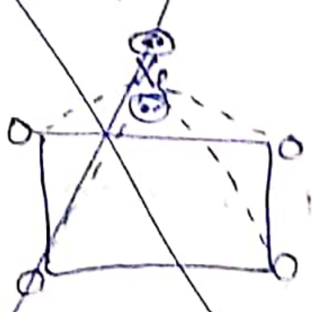
$sp^2 \times 3$

σ bond = 2

2π $2P_\pi - {}^3P_\pi = 0$ 1

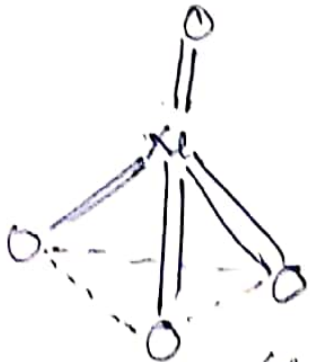
$2P_\pi - 3D_\pi = 0$ 1

~~③ XeO_4 (sp^3d^2)~~



~~Xe: [↑↓] [↑↓ | ↑↓ | ↑↓]~~

③ XeO_4 (sp^3)



Xe: [Kr] [↑↓] [↑↓ | ↑↓ | ↑↓] [↑↓ | ↑↓ | ↑↓ | ↑↓ | ↑↓] [| | | |]

Xe*: [↑] [↑ | ↑ | ↑] [↑ | ↑ | ↑ | ↑ |]

sp^3

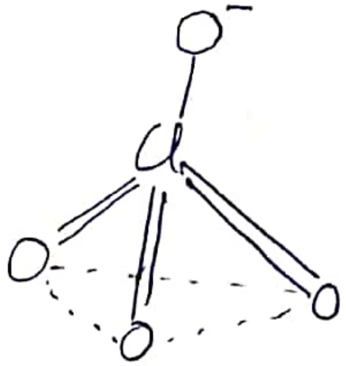
[↑ | ↑ | ↑ |]

4σ

($2P_\pi - 5D_\pi$) $P_\pi - P_\pi = 4$

$P_\pi - P_\pi = 0$

④ ClO_4^- (sp^3)



$p\pi - p\pi = 0$ ✓

$p\pi - d\pi = 3$ ✓

~~Cl~~

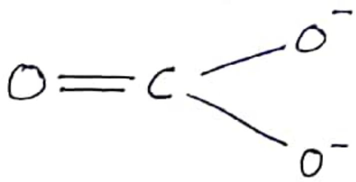
⑤ C_2H_2 (sp)

$p\pi - p\pi = 2$ ✓

$p\pi - d\pi = 0$ ✓

$H - C \equiv C - H$

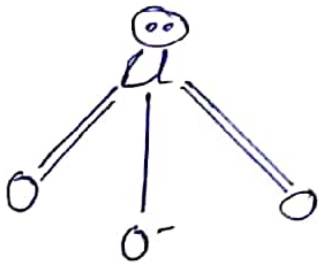
⑥ CO_3^{2-} (sp^2)



$p\pi - p\pi = 1$ ✓

$p\pi - d\pi = 0$ ✓

⑦ ClO_3^- (sp^3).

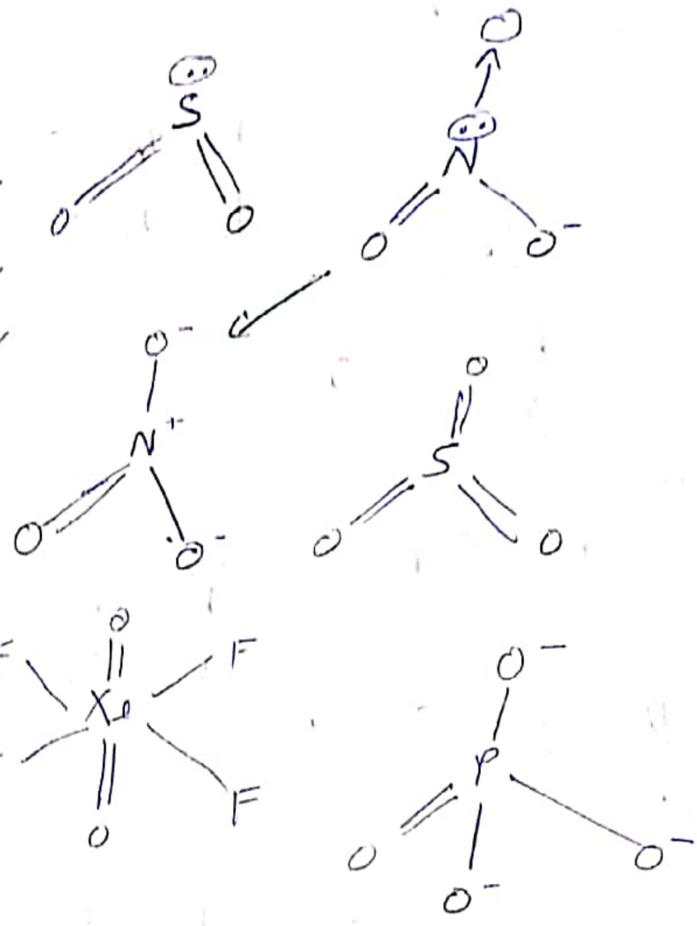


$p\pi - p\pi = 0$ ✓

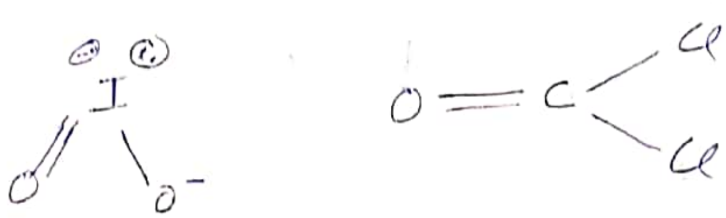
$p\pi - d\pi = 2$ ✓

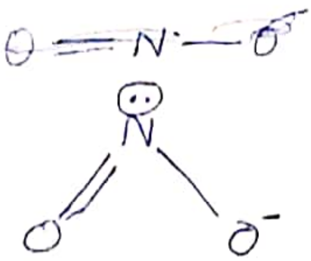
Q2.

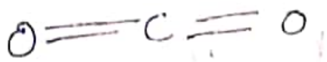
	$p\pi$-$p\pi$	$p\pi$-$d\pi$
SO_2 (sp^2)	1	1
NO_3^- (sp^2)	1	0
SO_3 (sp^2)	1	2 ✓
XeO_2F_4 (sp^3d^2)	0	2 ✓
PO_4^{3-} (sp^3)	0	1 ✓
IO_2^- (sp)	0	1 ✓
$COCl_2$ (sp^2)	1	0
NO_2^- (sp^2)	1	0
CO_2 (sp)	2	0
NOF_3 (sp^3)	0	0

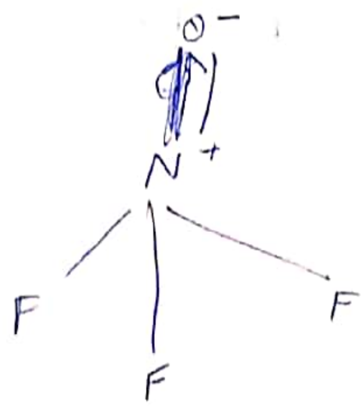


$\boxed{Q4}$









Solid/Liquid state hybridisation (Ratna Hai)

	Cationic Part	Anionic Part
(1) $2PCl_{5(s)}$	PCl_4^+	PCl_6^-
(2) $N_2O_{5(s)}$	NO_2^+	NO_3^-
(3) $N_2O_{4(s)}$	NO^+	NO_3^-
(4) $N_2O_{3(s)}$	NO^+	NO_2^-
(5) $Cl_2O_{6(s)}$	ClO_2^+	ClO_4^-
(6) $2BrF_{3(s)}$	BrF_2^+	BrF_4^-
(7) ~~$2IF_{(s)}$~~ $2IF_{5(s)}$	~~ICl_2~~ IF_4^+	IF_6^-
(8) $I_2Cl_{6(l)}$	ICl_2^+	ICl_4^-
(9) $PBr_{5(s)}$	PBr_4^+	Br^-
(10) $PI_{5(l)}$	PI_4^+	I^-
(11) $XeF_{6(s)}$	XeF_5^+	F^-
(12) $3I_2(l)$ or $3I_2(s)$	I_3^+	I_3^-

Q Comment on hybridisation & shape of cationic part of

I_2Cl_6 (l)

ICl_2^+

$\boxed{SP^3}$

~~ICl_4^-~~

~~SP^3D^2~~

$\boxed{\text{Bent Shape}}$

Q2. What is the bond angle diff between anionic part of solid I_2 & anionic part of N_2O_5

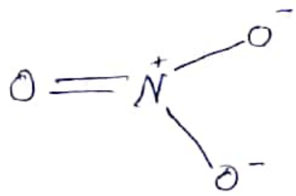
I_3^-

SP^3D

180°

NO_3^-

SP^2



120°

$180° - 120°$

$\boxed{60°}$

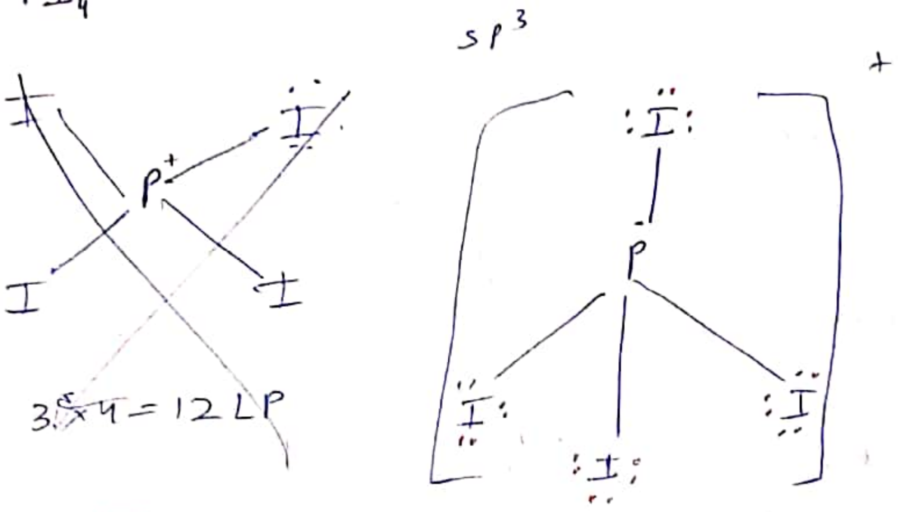
Q Comment hybridisation of anionic Part of $PBr_{5(s)}$

Br^-

No hybridisation

Q Comment Total no. of L.P on cationic Part of $PI_{5(s)}$

PI_4^+



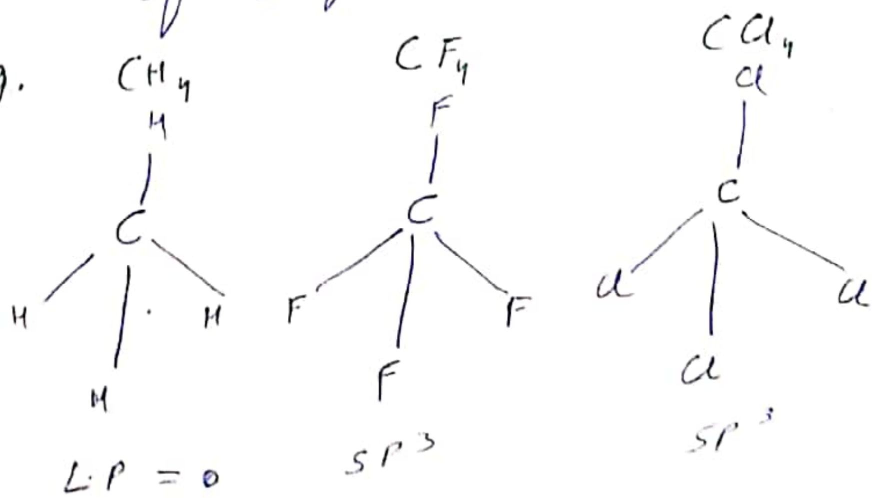
$3 \times 4 = 12$ LP

$\boxed{L.P = 12}$

Bond Parameter :-

① Bond Angle :-

Rule 1 → If hybridisation is same & There is no Lone Pair on central atom, Then angle remains same irrespective of size of atom. (Surrounding atoms must be ~~same~~ similar) same in to particular molecule)

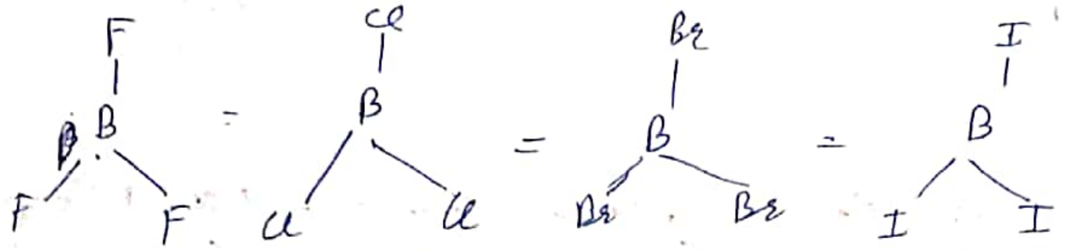
eg. 

$LP = 0$

sp^3

$BA = same = 109.5°$

Eg ② BF_3 BCl_3 BBr_3 BI_3



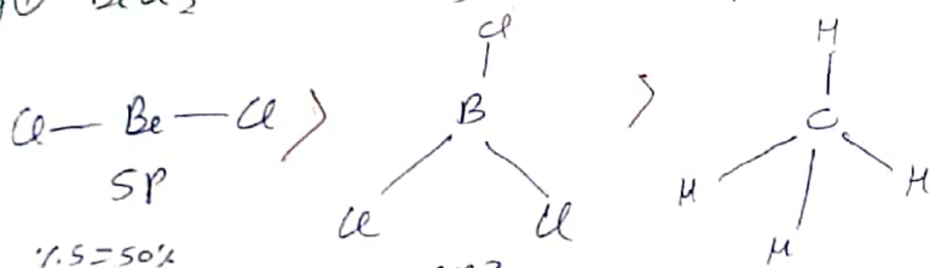
sp^2 sp^2 sp^2 sp^2

L.P = 0

Bond Angle = same = 120°

Rule 2 :- If hybridisation of central atom is Different then the bond angle will be directly proportional to % s character.

Eg ① $BeCl_2$ > BCl_3 > CH_4



SP sp^2

% S = 50% % S = 33.3 % S = 25%.

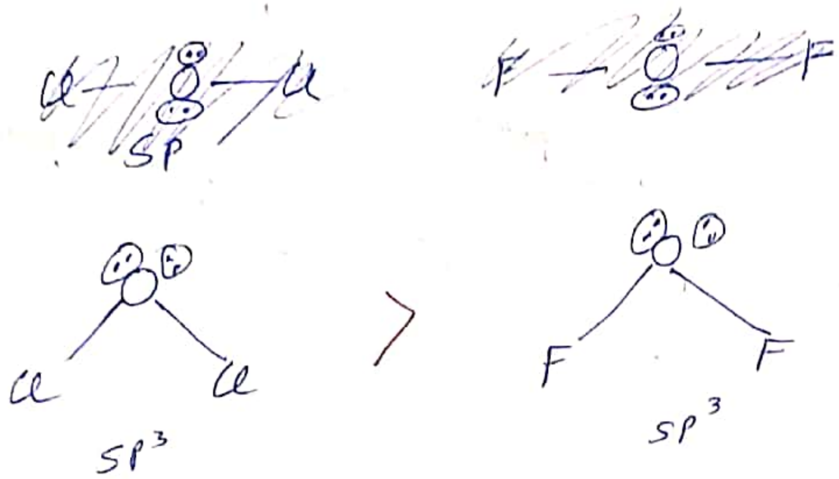
$\boxed{B.A \propto \% \text{ S character}}$

Rule 3.- If same type of hybridisation Takes place, then check steric repulsion or steric ~~hinderance~~. Steric crowding or steric hindrance.

$\boxed{\text{Size of surrounding atom} \uparrow \quad B.A \uparrow}$

Note - If central atom has at least one lone pair and surrounding atoms belongs to 3rd period elements onwards like Cl, Br or bulky group (CH_3, C_2H_5 etc) They steric crowding will appear.

Eg ① OCl_2 > OF_2



Size of Cl > size of F

BA(OCl_2) > BA(OF_2)

Eg ② OCl_2 < OBr_2

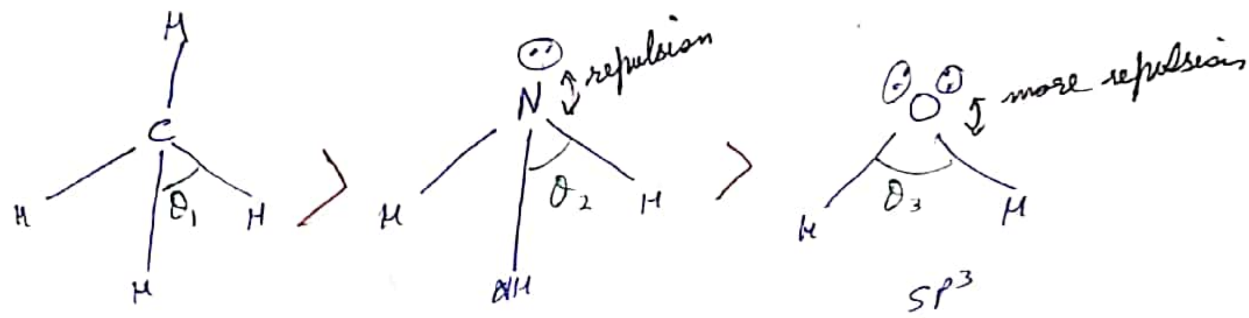
Size of Cl < Size of Br

BA(Cl < BA(Br)

Rule 4:- If same type of hybridisation takes place and there is no steric ~~...~~ crowding then check no. of lone pairs on central atom.

Note - In general no. of lone pairs on central atom increases bond angle decreases.

Eg ① CH_4 NH_3 H_2O



sp^3 sp^3

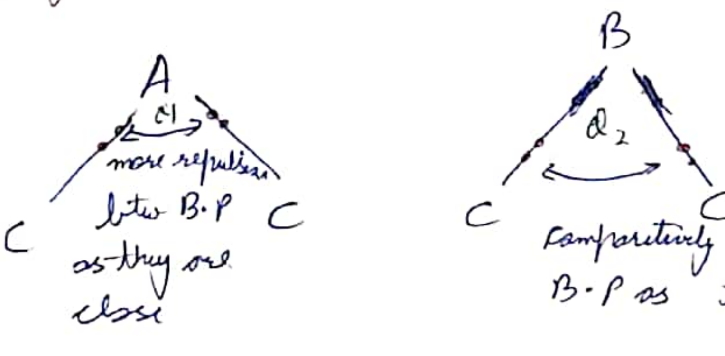
109.5° $<109.5^\circ$ $<<109.5^\circ$

$\theta_1 > \theta_2 > \theta_3$ $\boxed{\text{No. of L.P} \uparrow ;\ \text{B.A} \downarrow}$

Rule 5:- If all above factors are not working then check electron negativity of central atom or surrounding atom.

i) EN of Central Atom {All the surrounding atoms are same}

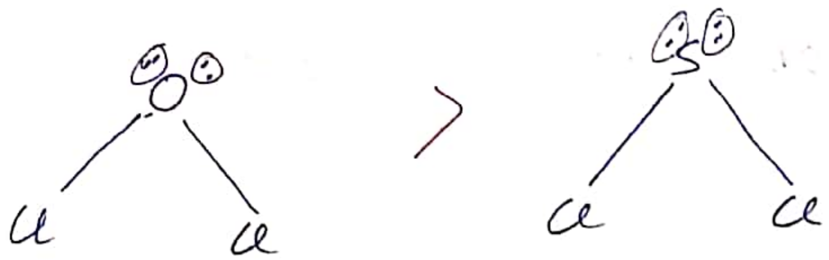
EN(A) > EN(B)



Comparatively less repulsion between B.P as they are far.

$\boxed{\theta_1 > \theta_2}$

Eg ①. OCl_2 SCl_2



$\boxed{\text{EN of central atom} \uparrow ;\ B\cdot A \uparrow}$

ii) E.N of surrounding atom {Same central atom}

$EN(B) > EN(C)$

$\boxed{\text{EN of surrounding atom} \uparrow ;\ B\cdot A \downarrow}$

Eg OH_2 > OF_2

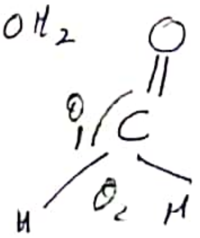
$EN(H) < EN(F)$

$BA\cdot(OH_2) >$ ~~OH~~ $B\cdot A(OF_2)$

Rule 6 - Multiple Bond repel more strongly as compared to single bond.

$(PB - P\cdot B) > (PB - SB) > (SB - SB)$

Eg ① COH_2



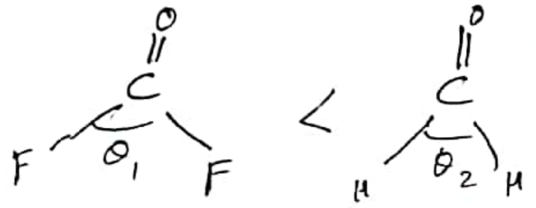
$\boxed{\theta_1 > \theta_2}$

Q Bond Angle order?

① $OF_2 < OCl_2 < OBr_2 < OI_2$ Steric Crowding

$OI_2 > OBr_2 > OCl_2 > OF_2$

②



$\theta_1 < \theta_2$ EN of surrounding atom

$EN(F) > EN(H)$

③ $NH_2^- < NH_3 < NH_4^+$

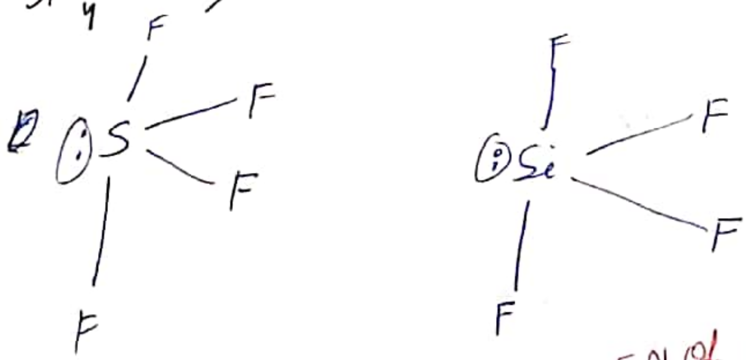
sp^3 sp^3 sp^3

~~Diff Hybridisation~~

$LP = 2$ $LP = 1$ $LP = 0$

Lone Pair

④ $SF_4 > SeF_4$



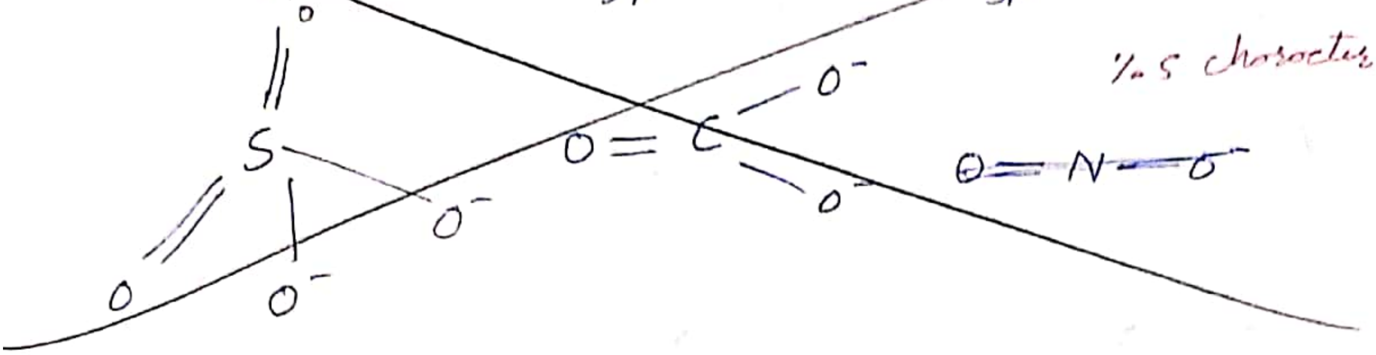
EN of central atom

$EN(S) > EN(Se)$

~~Q.5. $SO_4^{2-} < CO_3^{2-} < NO_3^-$~~

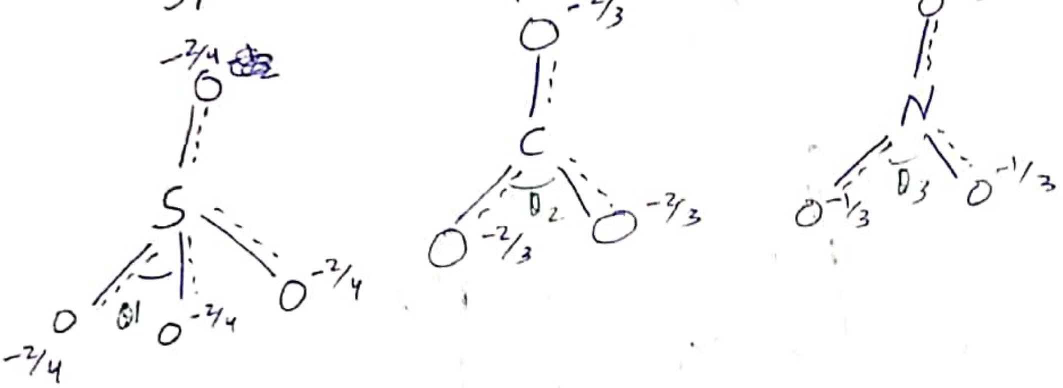
~~sp^3 sp^2 sp^2~~

~~% s character~~



Q(5) SO_4^{2-} CO_3^{2-} NO_3^-

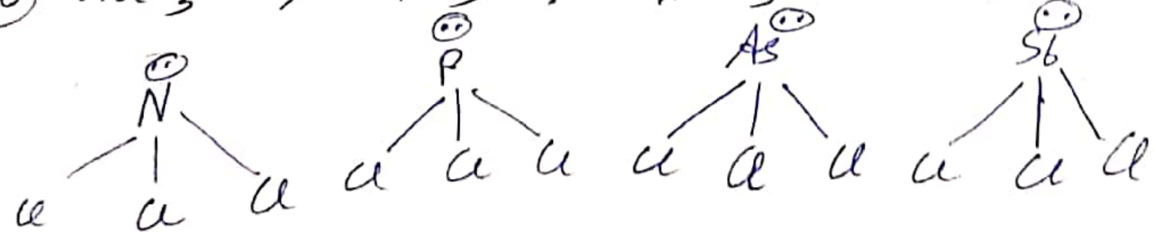
sp^3 sp^2 sp^2



$\theta_1 < \theta_2 = \theta_3$ {B.A $\propto$ % s character}

~~[illegible]~~

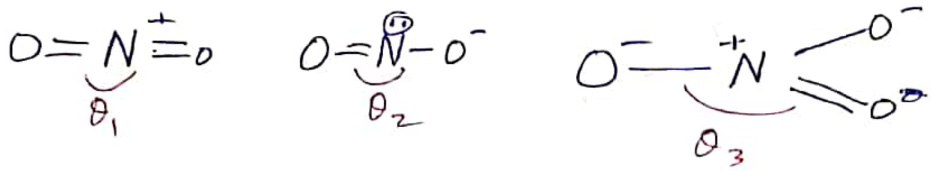
(6) $NCl_3 > PCl_3 > AsCl_3 > SbCl_3$



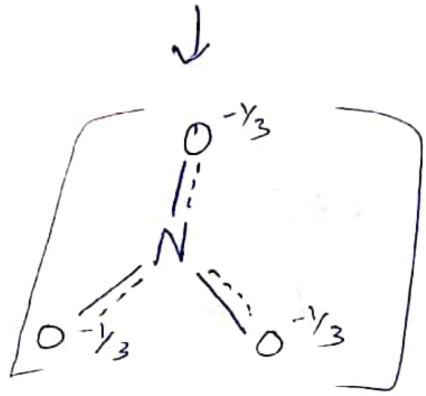
EN of central atom.

7 NO_2^+ NO_2^- NO_3^-

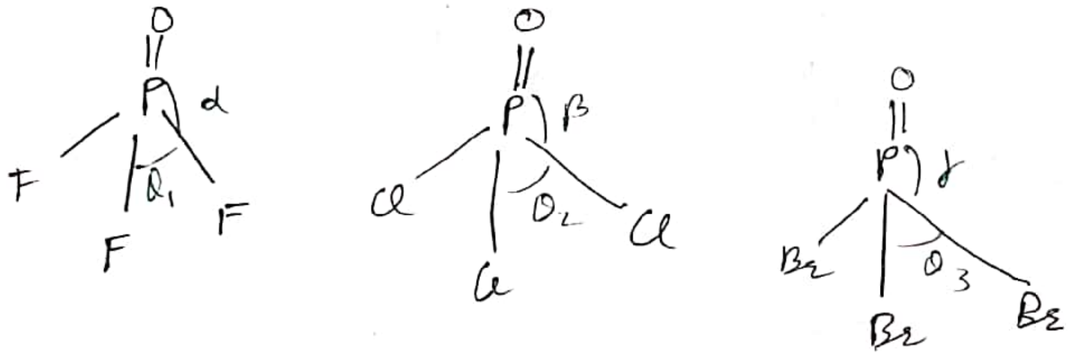
sp sp^2 sp^2



$\boxed{\theta_1 > \theta_3 > \theta_2}$



8



$\theta_3 > \theta_2 > \theta_1$ {steric crowding}

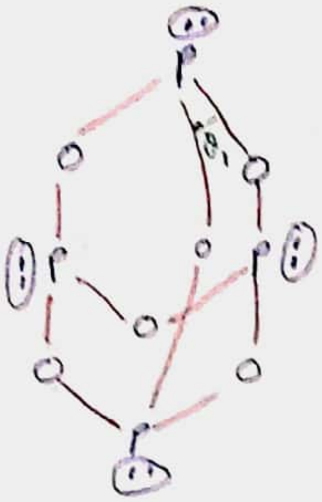
$\alpha > \beta > \gamma$ {depends on θ}

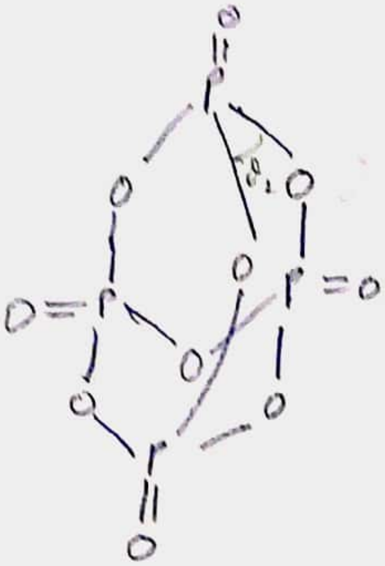
9 H_2O < $H_3O^{\oplus}$ (Lone Pair)

10 SO_2 < SO_3 (Lone Pair)

11

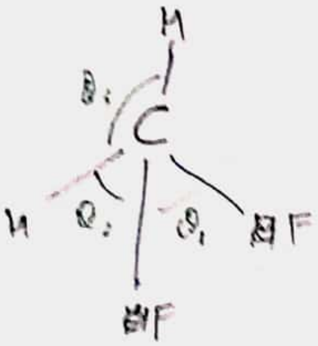
(11) P_4O_6 $<$ P_4O_{10} (Lone Pair)





$L.P_{Repulsion} > (B.P)_{Repulsion}$

(12) CH_2F_2



$\theta_3 > \theta_2 > \theta_1$ (E.N, repulsion)

(13) H_3O^+ NH_4^+

(14) CO_2 $AlCl_3$ CH_4

(15) NH_4^+ NH_3

(16) NH_4^+ NH_3 H_2O

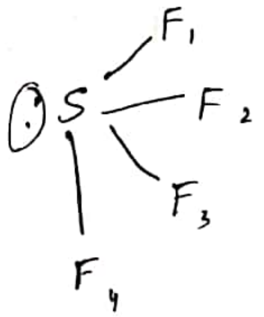
(17) OH_2 OF_2 OCl_2 OBr_2

(18) NH_3 NF_3 NCl_3

(19) OH_2 $O(CH_3)_2$

Q(20) $BeCl_2$ BF_3 CH_4

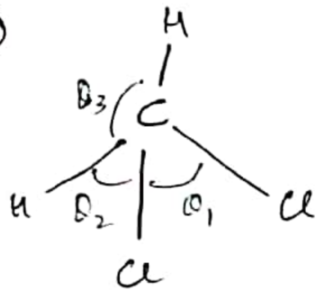
Q(21).

(i) 

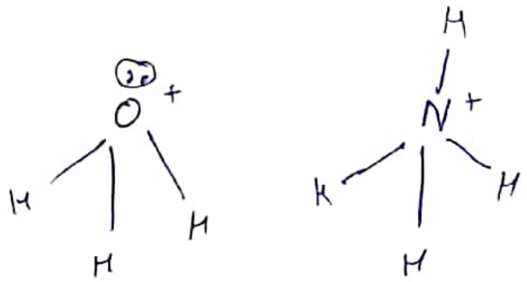
$F_1 \widehat{S} F_4$

$F_2 \widehat{S} F_3$

(22)



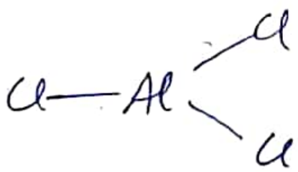
(13) $H_3O^+ < NH_4^+$ (Lone Pair)



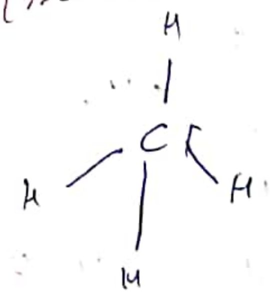
(14) $CO_2 > AlCl_3 > CH_4$ (% s character)

$O=C=O$

sp

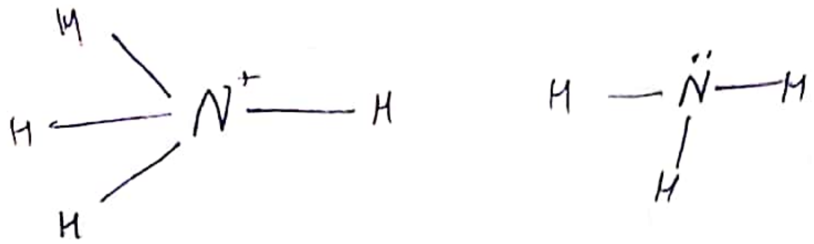


sp^2

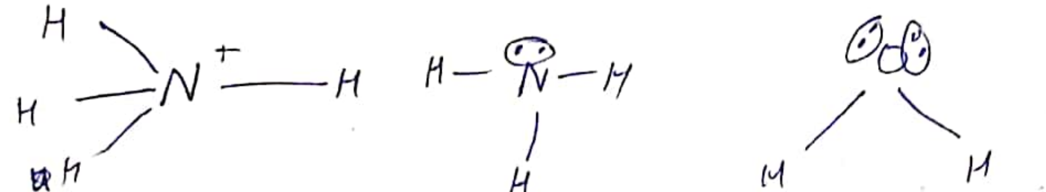


sp^3

(15) $NH_4^+ > NH_3$ (Lone Pair)



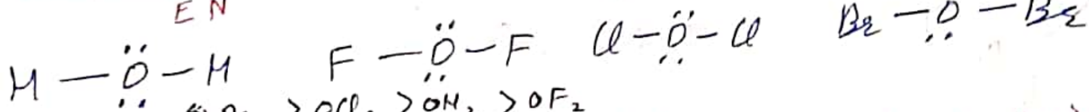
(16) $NH_4^+ > NH_3 > H_2O$ (Lone Pair)



sp^3 sp^3 sp^3

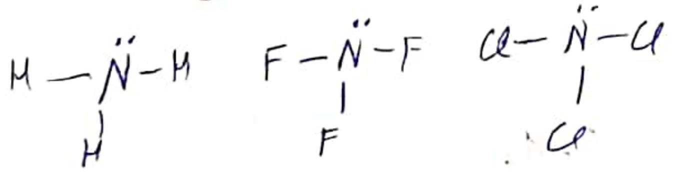
EN_9

(17) $\underbrace{OH_2 > OF_2}_{EN}$, $\underbrace{OCl_2 < OBr_2}_{\text{steric crawding}}$, (steric crawding)



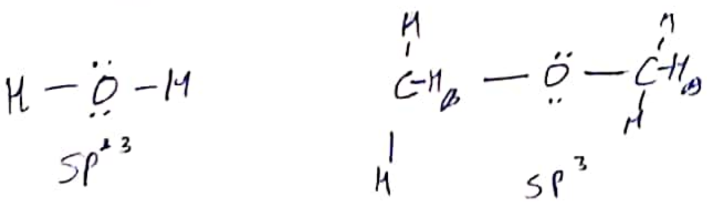
$OBr_2 > OCl_2 > OH_2 > OF_2$

(18) $\underbrace{NH_3 > NF_3}_{EN}$, NCl_3 (steric crawding $\&_9$ EN)



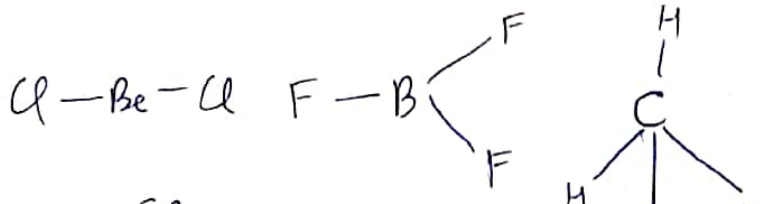
$NCl_3 > NH_3 > NF_3$

(19) $OH_2 < O(CH_3)_2$ (steric crawding)



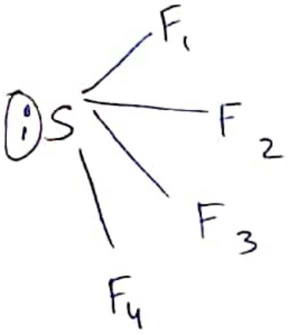
sp^3 sp^3

(20) $BeCl_2$ > BF_3 > CH_4 (% s character)



sp $\qquad$ sp^2 $\qquad$ sp^3

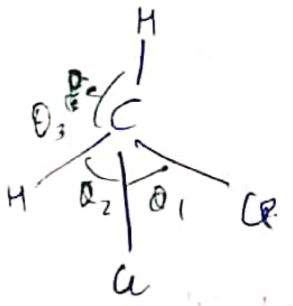
(21)



$\widehat{F_1SF_4}$ > $\widehat{F_2SF_3}$

(<180°) $\qquad$ (<120°)

(22)

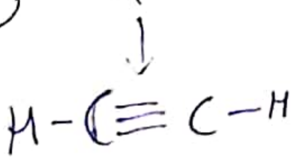


~~$\theta_1 > \theta_2 > \theta_3$~~ ~~(EN)~~

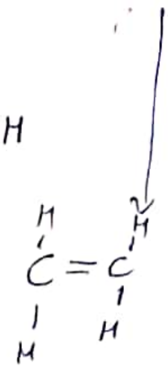
~~(Steric crowding)~~

$\theta_1 > \theta_3 > \theta_2$

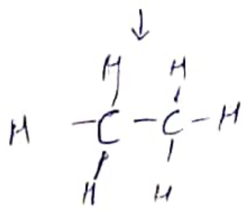
(23) C_2H_2 > C_2H_4 > C_2H_6 (Hybridisation)



sp



sp^2



sp^3

(24) $NF_3 < NCl_3 < NBr_3 < NI_3$ (steric crowding)

(25) $O(CH_3)_2 < OCl_2 < OBr_2 < OI_2$ (Exception)

Drago's Rule :-

hybridisation fail

B.A. $NH_3 > PH_3 > AsH_3 > SbH_3$

$H_2O > H_2S > H_2Se > H_2Te$

$PH_3 \rightarrow 94°$
$AsH_3 \rightarrow 92.5$
$SbH_3 \rightarrow 91°$

$H_2S \rightarrow 92°$
$H_2Se \rightarrow 91°$
$H_2Te \rightarrow 90.5°$

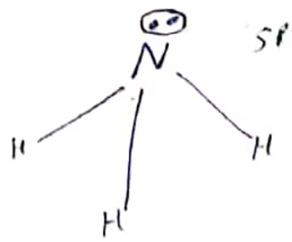
Angles are around ~~more than~~ 90° means pure orbitals are participating rather than hybrid orbitals.

→ According to this rule, if central atom belongs to 15th or 16th group of 3rd period element onwards and e^- negativity of surrounding atom is ≤ 2.1. Then in that compound bond angle observed is almost 90°. which indicates that no or negligible hybridisation takes place & bond is formed by almost pure p-atomic orbitals.

Explanation ① - $\ddot{N}H_3$

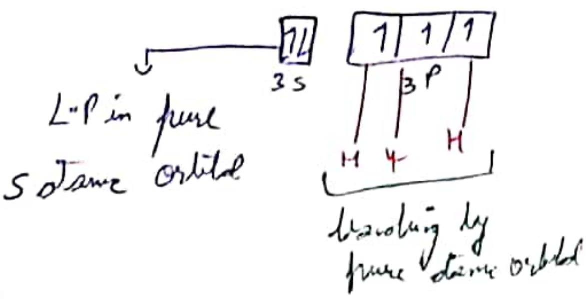
N: [He] $2s^2$ $2p^3$

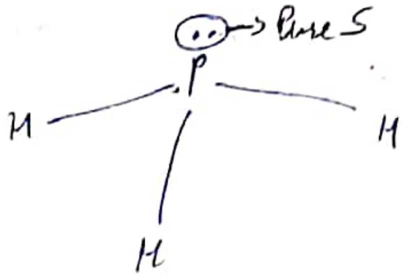
sp^3



$\ddot{P}H_3$

$P: [Ne]\ 3s^2\ 3p^3$





Explanation-2:- The cause of no hybridisation in this molecule can be explained as follows.

→ In case of ~~P~~ Phosphorous, the energy required in hybridisation is about 600 kJ/mol. which is not compensated by energy released from the bond formation using hybrid orbitals.

Q1. ~~Ammonia~~ Ammonia is strong Lewis base as compared to Phosphine (PH_3)

Q2. Formation of Ammonium ion is much easier ~~the~~ than ~~Phosphonium~~ Phosphonium ~~ion~~ ion (PH_4^+) why?.

Q.3. Ammonia is easily soluble in water than PH_3 why?

→ sp^3 orbital ⇒ 25% s character ⇒ Attraction from nucleus is less

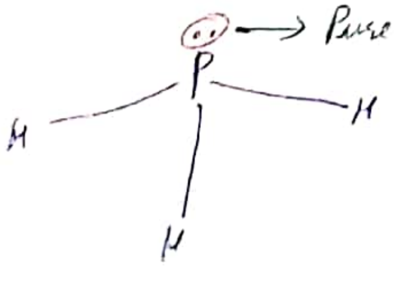
↓

e^- remove easily

→ Pure s orbital ⇒ 100% s character ⇒ Attraction of nucleus is more

↓

removing e^- require more energy.



Q1. A lewis base ~~to~~ donate ~~o~~ lone pair easily & as NH_3 L.P require very less energy to remove lone pair than PH_3. Thus, ~~A~~ It is stronger lewis base.

Q2. $\ddot{N}H_3 + H^{\oplus} \xrightarrow{\text{Room Temp}} NH_4^{\oplus}$ (Donates L.P easily)

$\ddot{P}H_3 + H^{\oplus} \xrightarrow{>800°C} PH_4^{+}$

Q3. $NH_3 + \underset{\substack{\downarrow \\ H^+ \; OH^-}}{H_2O} \longrightarrow NH_4^{\oplus} \; OH^{\ominus}$

→ In ammonia, L.P of Nitrogen is present in one of the SP^3 hybridised orbital which is more directional & more reactive.

→ In PH_3 (Phosphene), L.P of Phosphorous is present in almost pure S-orbital which is Less directional or less reactive.

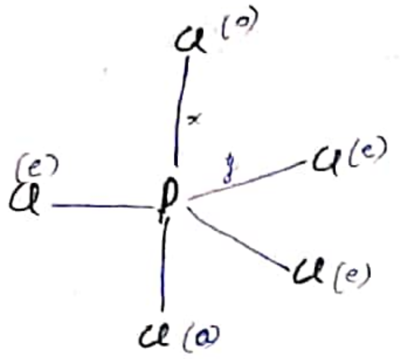
Types of Hybridisation

Equivalent Hybridisation
→ no L.P on central atom
→ Surrounding atoms must be same
→ All bond angles are same
Eg CH_4, CCl_4, BCl_3

Non Equivalent Hybridisation
→ If there is L.P on central atom, surrounding atoms are not same or different Bond angles
Eg NH_3, $COCl_2$, PCl_5, IF_7, XeF_4

Bents Rule :-

→ It is applicable for non - equivalent hybridisation.



$\underbrace{S + P_x + P_y}_{\text{forms hybrid orbital}} + \underbrace{P_z + d_{z^2}}_{\text{forms hybrid orbital differently}}$

$\boxed{d_{(P-Cl)_{axial}} > d_{(P-Cl)_{eq.}}}$ → Bond length d = Bond Length.

$x > y$

→ In $PCl_{5(g)}$ It is observed that there are two longer (weaker) axial bonds and 3 shorter (stronger) equitorial bonds are present.

→ Experimentally It is concluded that all hybrid orbitals are not equal In PCl_5.

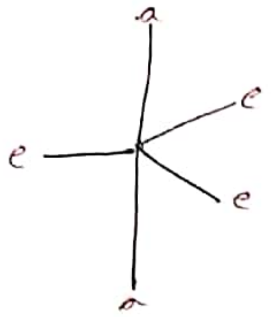
→ Equitorial bonds are formed by SP^2 hybrid orbitals and axial bonds are formed by ~~Pd~~ PD hydbrid orbitals of central atom

According to Bents rule.

→ More Electron Negative element (surrounding atom) prefer to stay in the hybrid orbital of central atom. which is having less % s character and is far from nucleus.

→ while L.P prefer to stay in the hybrid orbital of central atom which is having more % s character or is closer to ~~the~~ nucleus.

In SP^3D hybridisation (TBP)



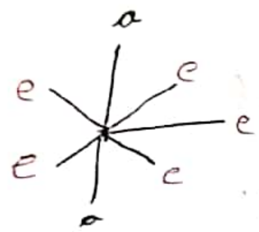
$\boxed{d_{axial} > d_{equitorial}}$

$\underbrace{S + P_x + P_y}_{SP^2} + \underbrace{P_z + d_{z^2}}_{PD}$

PD → directional nature ↑; P% ↑
S% ↓ ⇒ more EN will stay there
length of H.O ↑

SP^2 → directional nature ↓; %P ↓ length of H.O ↓
%S ↑ L.P or double will stay.

In SP^3D^3 hybridisation (PBP)



$\underbrace{S + P_x + P_y + D_{xy} + D_{x^2-y^2}}_{SP^2P^2} + \underbrace{P_z + PD_{z^2}}_{PD}$

SP^2P^2 → directional nature ↑;
%P ↑; length of H.O ↑
%S ↓; more EN will stay

$\boxed{d_{equitorial} > d_{axial}}$

PD → Directional nature ↓;
%P ↓; length of H.O ↓
%S ↑; L.P & Double bond will stay.

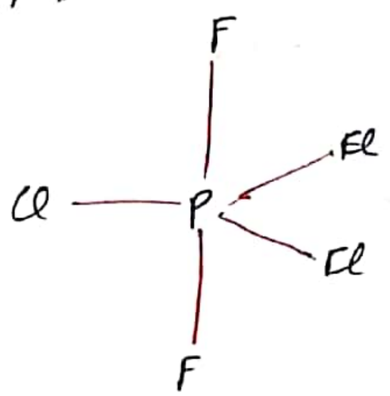
Note:-

1. SP^3D – A strong electron Negative element will prefer axial position in sp^3d hybridisation.
2. A L.P or a double bond occupy equitorial position in SP^3D hybridisation because they want to be close to the nucleus.

2. SP^3D^3 - Strong EN element will prefer equitorial position in ~~SP~~ SP^3D^3 hybridisation and L.P or double bond pair will prefer axial position.

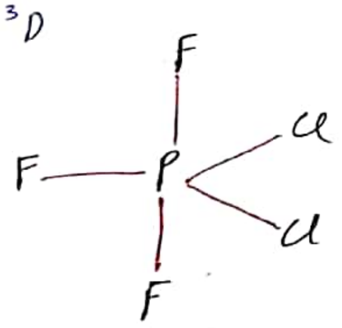
① PCl_3F_2

SP^3D



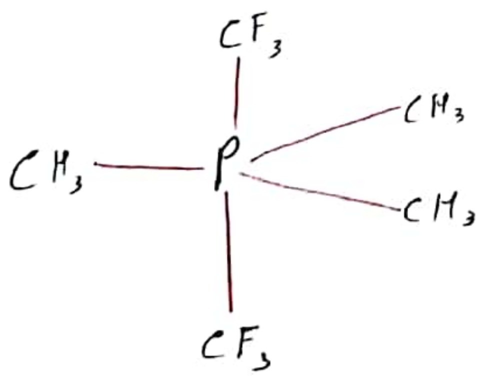
② PF_3Cl_2

SP^3D

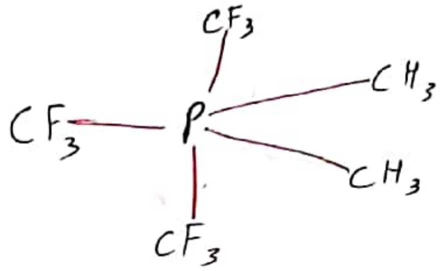


③ $P(CH_3)_3(CF_3)_2$

SP^3D

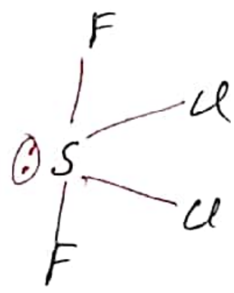


④ $P(CH_3)_2(CF_3)_3$ SP^3D

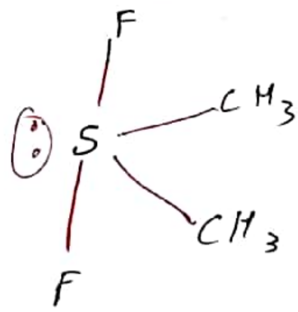


⑤ SF_2Cl_2

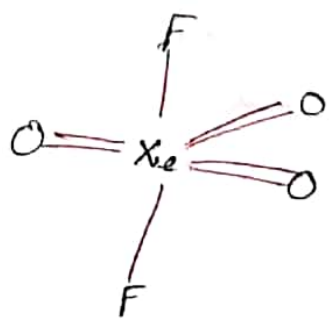
SP^3D



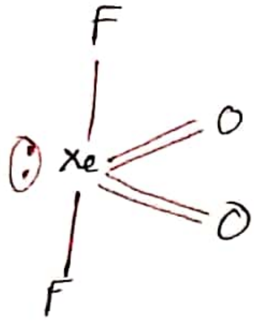
⑥ $SF_2(CH_3)_2$



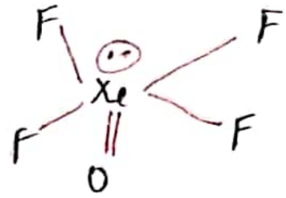
⑦ XeO_3F_2



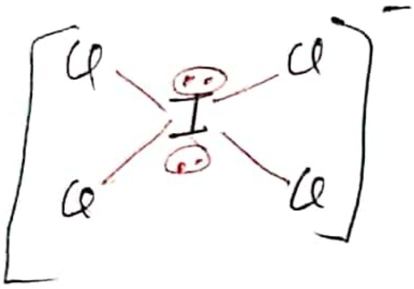
(8) XeO_2F_2



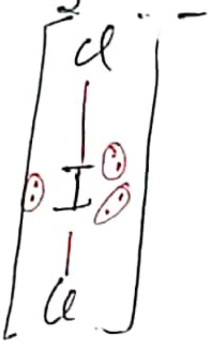
(9) $XeOF_4$



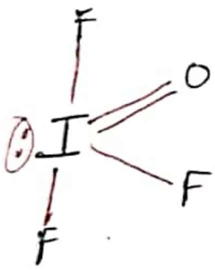
(10) ~~ICl_4~~ $ICl_4^{\ominus}$



(11) $ICl_2^{\ominus}$



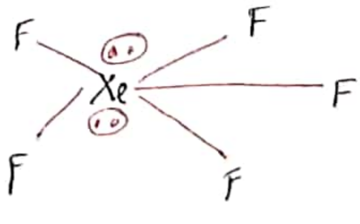
(12) IOF_3



(13) XeF_2

(14) IO_3^-

(15) XeF_5^- sp^3d^3

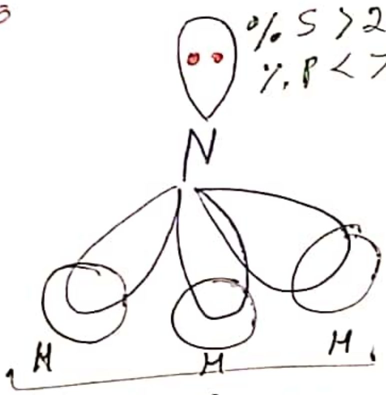


Alternative Statements of Bents Rule:-

→ More E.N atom not only prefer to stay In the hybrid orbital of central atom which is having less % s character but also It decreases % s character to attached hybrid orbital of central atom which depends upon circumstances.

→ L.P. not only to prefer to stay in the hybrid orbital of central atom which is having more % s character but also it increases % s character to attached hybrid orbital of central atom which depends upon circumstances.

NH_3



% S > 25%
% P < 75%

% S < 25
% P > 25

जहा L:P जाएगा, % S charactr बडा देगा

जहा EN जाएगा % s charactr कम कर देगा

How To solve Questions of Bond Length.

Rule I - If shell no. of central atom or surrounding atom are different.

$\boxed{\text{Shell no.} \uparrow \quad \text{Bond Length} \uparrow}$

eg. H - F, H - Cl, H - Br, H - I

$d = (R_H + R_F)$ $(R_H + R_{Cl})$ $(R_H + R_{Br})$ $(R_H + R_I)$

$H-F < H-Cl < H-Br < H-I$

Eg. $d_{(P-O)} > d_{(S-O)} > d_{(Cl-O)}$

$r_P > r_S > r_{Cl}$

Eg ③. $d_{(P-F)} < d_{(P-Cl)}$

Eg ④. $d_{(N-H)} (NH_3) < d_{P-H} (PH_3)$

Rule-② – If surrounding atoms are same, central atoms are different & central atoms belong to same period.

left Zeff ↑ Size ↓ Right

Bond length ↓

Eg ① $d_{C-O} > d_{N-O}$

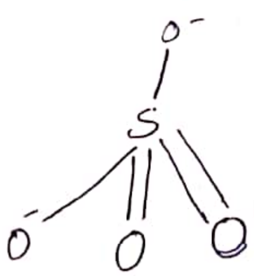
Rule ③ → If central atoms & surrounding atoms are same

$\boxed{B.O \uparrow ;\ B.S \uparrow ;\ B.L \downarrow}$

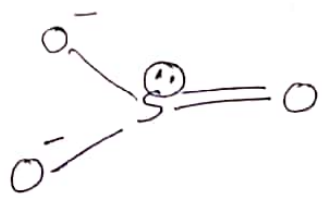
Eg ① $d_{(C-O)} (CO_2) < d_{C-O} (CO_3^{2-})$

Eg ② $d_{S-O} (SO_4^{2-}) < d_{S-O} (SO_3^{2-})$

SP_3

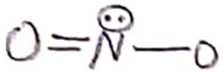


$BO = \frac{6}{4}$

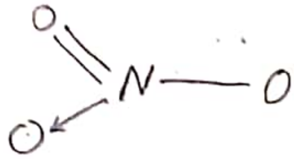


$B.O = \frac{4}{3}$

Eg. $d_{N-O}(NO_2^-) < d_{N-O}(NO_3^-)$



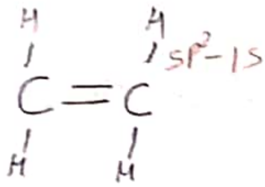
$B \cdot O = \frac{3}{2}$

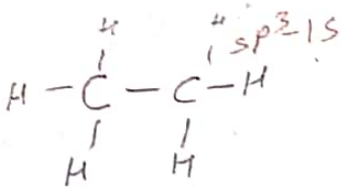


$B \cdot O = \frac{4}{3}$

Rule-4 → If central atom, surrounding atom and Bond order is same Then Bond Length depends on hybridisation

eg 1. $C_2H_4 < C_2H_6$





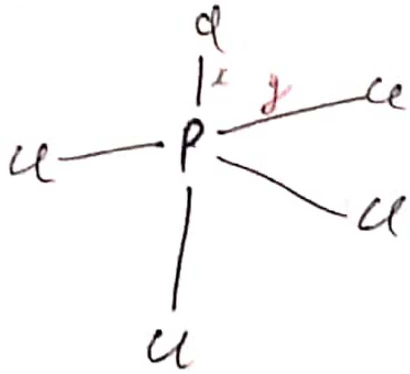
sp^2 sp^3

$\boxed{\%S \uparrow \quad B \cdot L \downarrow}$

$d_{(C-H)}\ C_2H_6 > d_{(C-H)}\ C_2H_4$

Rule-5 → If central atom, surrounding atom, Bond order and hybridisation are same then Bond length is decided by Bents rule.

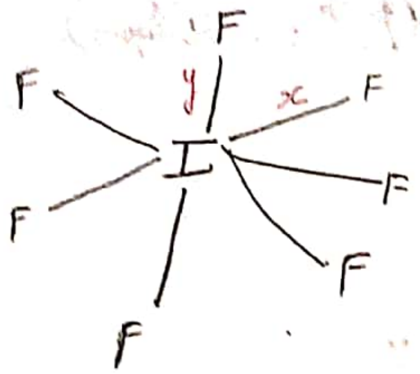
Eg 1



$d_{(P-Cl)\,axial} > d_{(P-Cl)\,equitorial}$

$x > y$

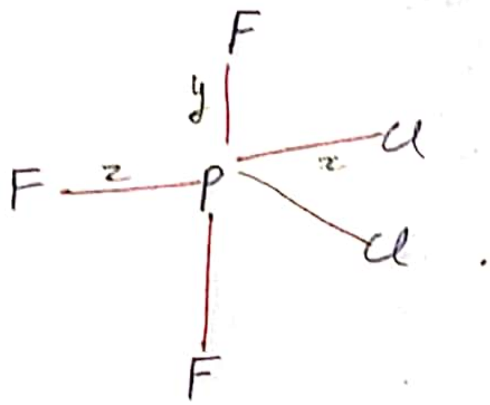
eg ②



$d_{(I-F)}$ equitorial $> d_{(I-F)}$ axial

$x > y$

Q PF_3Cl_2



~~y > z~~
$x >> z$

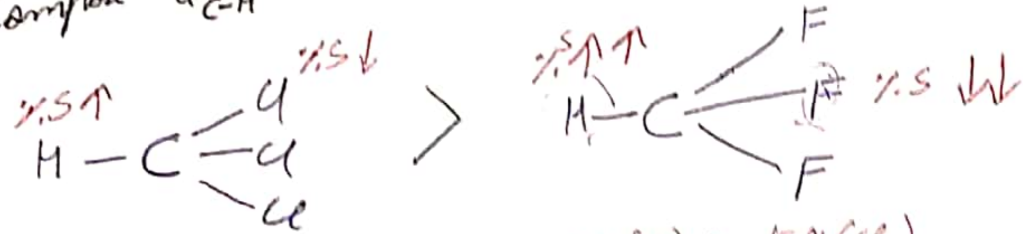
i) $d_{P-F}(e) < d_{P-F}(a)$ $z < y$

ii) $d_{P-F}(e) < d_{P-Cl}(eq)$ $z < x$

iii) $d_{P-F}(a) < d_{P-Cl}(eq)$

Bent Gaya Bhad me
(Size is always dominating)

Q Compare d_{C-H} in $CHCl_3$ & CHF_3



$EN(F) > EN(Cl)$

CHF_3 में %S ज्यादा बढ़ेगा, B.L कम होगी।

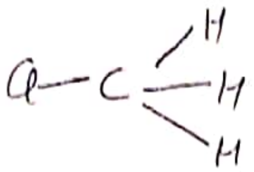
Q2. Compare C–X in $CHCl_3$ & CHF_3 (X → Halogen)

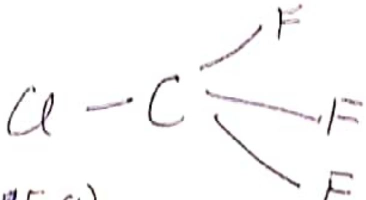
$CHCl_3 > CHF_3$

↓

Size ↑ B.L ↑

Q3. Compare d_{C-Cl} in CH_3Cl & CF_3Cl





$d_{(C-Cl)}^{(CH_3Cl)} > d_{(C-Cl)}^{(CF_3Cl)}$