

Chemical Bonding

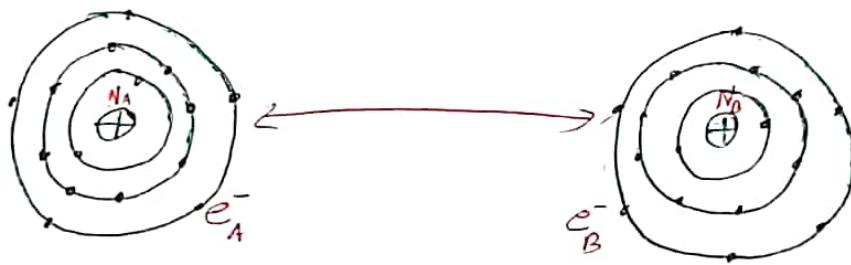
Chemical bond - It is the force of attraction between two atoms which binds them together.

Why a chemical bond is formed?

- To gain stability by decreasing their potential energy or to acquire inert gas configuration.
- Atoms are more reactive than molecules because they have higher energy than molecules.

$\text{Cl} \rightarrow$ atom (more reactive)

$\text{Cl}_2 \rightarrow$ molecule (less reactive)



Attraction

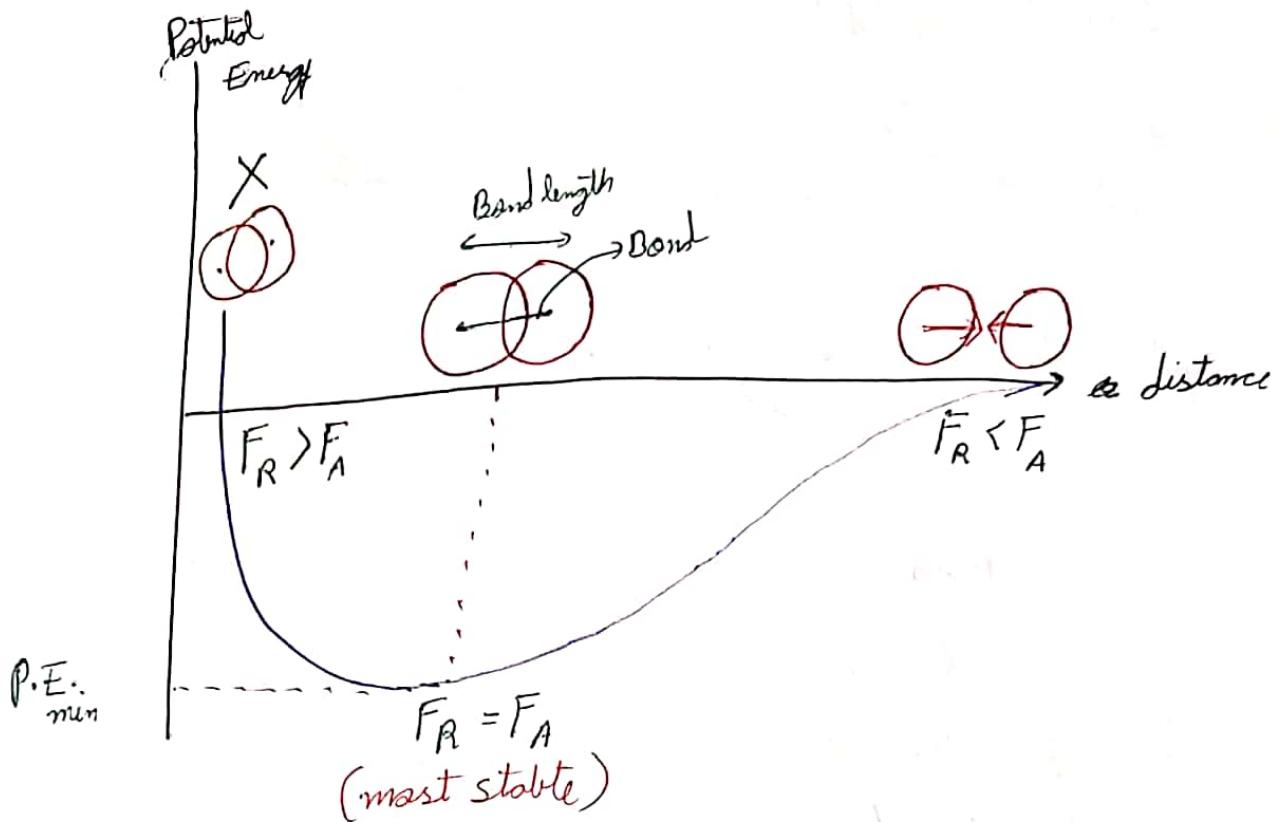
$$N_A - e_B$$

$$N_B - e_A$$

Repulsion

$$e_A - e_B$$

$$N_A - N_B$$



F_A = Force of Attraction

F_R = Force of Repulsion

- When atoms approaches to each other, attraction force dominates over repulsion.
- Bond formation takes place when attraction force is equal to repulsion force because at this point, potential energy is minimum & stability is max.

Bond Energy

- It is the amount of energy required to break bond.
- It is the amount of energy released during bond formation.

Bond Length (r_0/d)

→ It is the shortest internuclear distance where bond formation takes place between two atoms.

Types of Chemical Bonds.

Strong Bond

($E_{\text{Released}} > 42 \text{ kJ/mol}$)

① Ionic / Electrovalent Bond

② Covalent Bond

③ Co-ordinate Bond / Dotive Bond

④ Metallic Bond.

Weak Bond

($E_{\text{Released}} < 42 \text{ kJ/mol}$)

Hydrogen Bond

(~~E_{Released}~~ & $E_{\text{Rel}} < 42$)

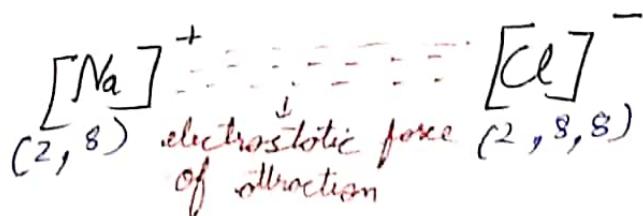
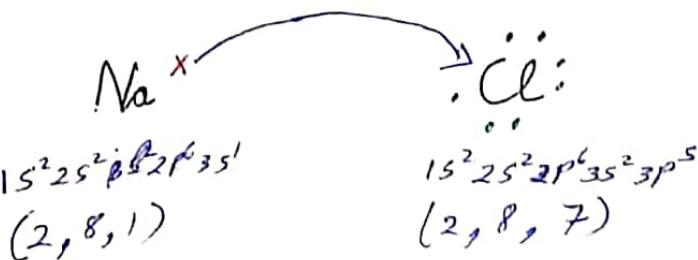
Weak forces

($E_{\text{Rel}} < 8 \text{ kJ/mol}$)

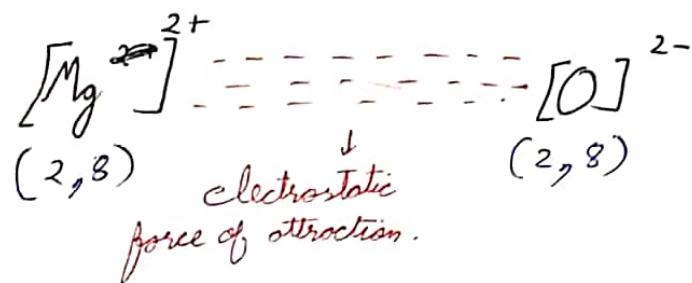
1) Ionic / Electrovalent Bond :-

→ It is the electrostatic force of attraction between two oppositely charged ions which are created by complete transfer of electron.

→ Eg:-



Eg 2.



XX → Ionic bond is non-directional in Nature.

Electrovalency → It is defined only for ionic bonds.

→ It is the number of electron transfer during formation of ionic bond by an atom.

Eg 1. NaCl E.v. (Na) = 1
E.v. (Cl) = 4

Ex 2. MgO E.v. (O) = 2
E.v. (Mg) = 2

3. AlF₃ E.v. (Al) = 3
E.v. (F) = 1



Condition for the formation of ionic bond →

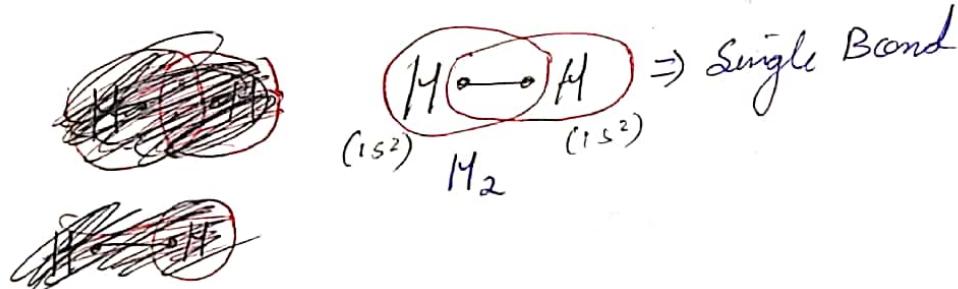
$$\rightarrow \boxed{\Delta E \cdot N \geq 2 \cdot 1}$$

- Low ionisation energy of electropositive element.
- High Electron Affinity of electronegative element.
- High Lattice energy.

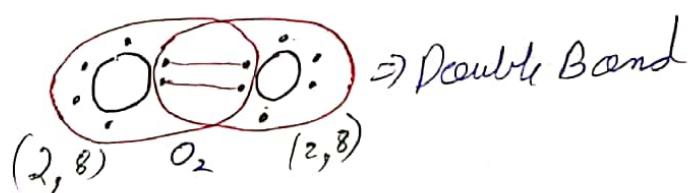
2. Covalent Bond:-

- This type of Bond is formed by equal sharing of electrons between two bonded atoms.
- ~~Covalent Bonds are also~~
- Covalent Bonds are directional in nature.

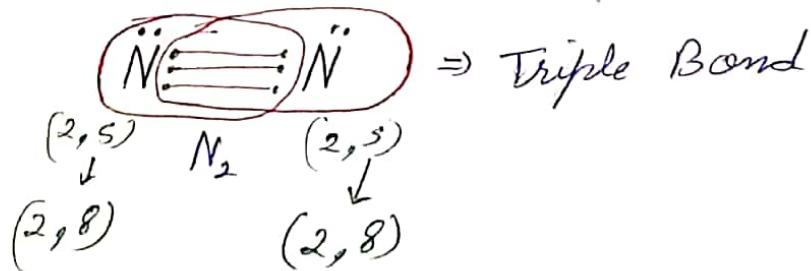
→ Eg 1.



Eg 2.



Eg 3.



Electron Pair

→ Pair of e^- is known as electron pair

Electron Pairs

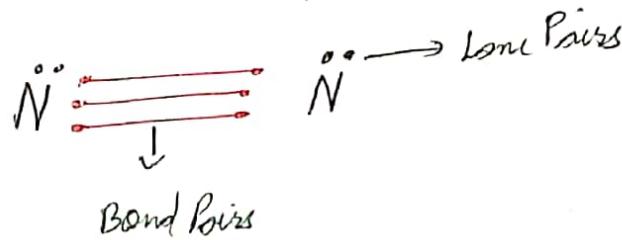
Bond Pair (B.P)

→ Pairs of e^- which are participating during bond formation

Lone Pair (L.P)

→ Pairs of e^- which do not participating during bond formation.

Eg1.



Bond Pairs - 3

Lone Pairs - 2

Q Count Bond Pairs Lone Pairs in

- ① Cl_2
- ② H_2O
- ③ CS_2
- ④ SF_2
- ⑤ HCl
- ⑥ NCl_3
- ⑦ C_2F_2
- ⑧ CO_2
- ⑨ NH_3
- ⑩ CH_4

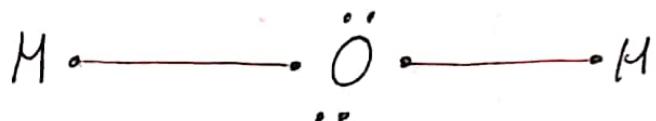
①



Bond Pair $\rightarrow 1$

Lone Pair $\rightarrow 6$

②



BP - 2

LP - 2

③

BP - 1
LP - 3



④



BP - 0 4

LP - 0 4

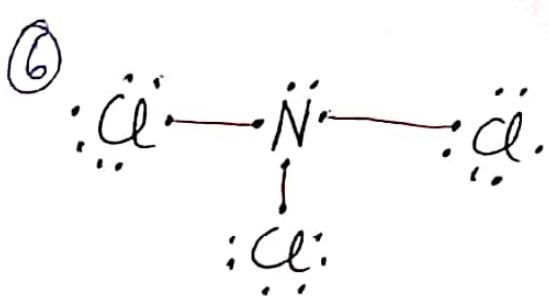
⑤



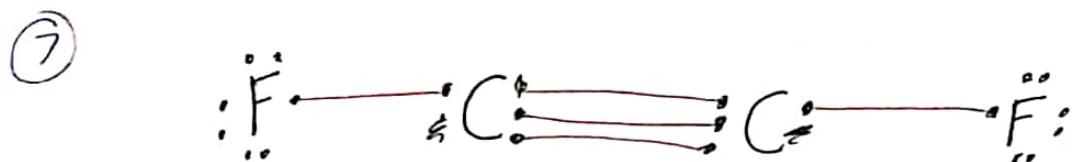
BP - 2

LP - 2 8

(180)



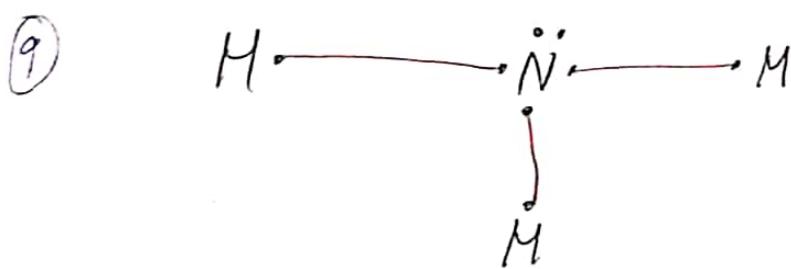
BP - 3
LP - 10



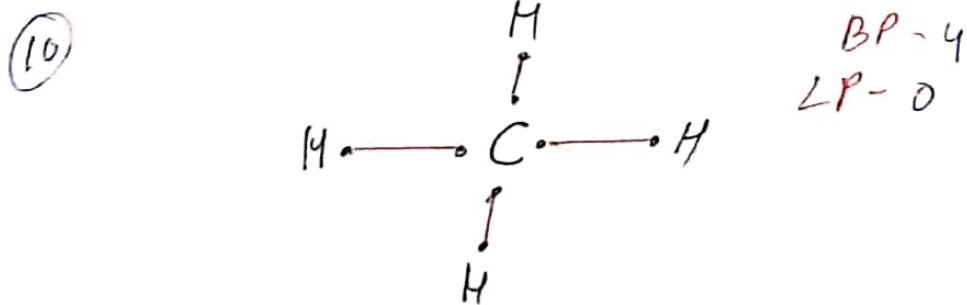
BP - 5
LP - 6



BP - 4
LP - 4



BP - 3
LP - 1



BP - 4
LP - 0

CovaLency - The number of electrons ^{not} participating of an element in p covalent bond formation is known as its covalency.

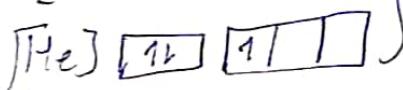
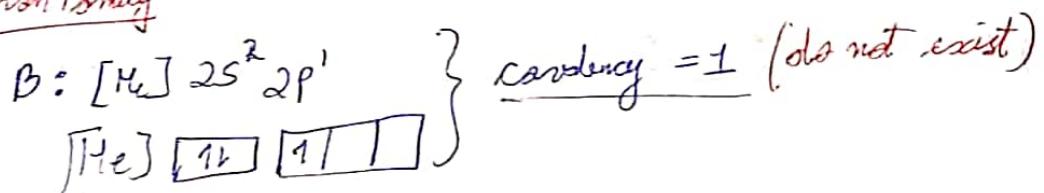
It is defined as the number of ^{or} covalent bonds formed by an atom in a covalent molecule or compound.

Number of unpaired electrons in outermost shell in ground state or in excited state decides co-valency of an atom.

→ In S-block & P-block elements, the number of e^- in outermost shell is called valence electron.

→ In D-block elements, outermost 's' & penultimate 'd' electrons are collectively called valence electrons.

Boron Family



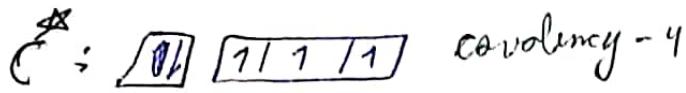
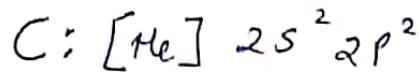
CovaLency = 3

→ The excitation of $1s$ is not possible since shell changes require very high amount of energy.

→ Boron, Aluminium, Indium, Gallium forms bond only in excited state and their co-valency is always ≥ 3 .

e.g. BCl_3 , BF_3 , $AlCl_3$ (covalent, not ionic bond) $\Delta EN < 2.1$

Carbon family - (tetravalent)

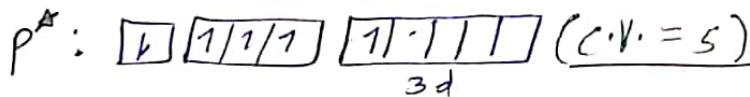
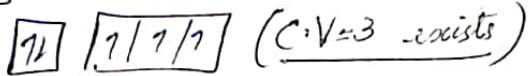
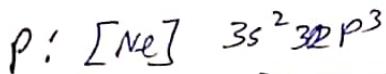
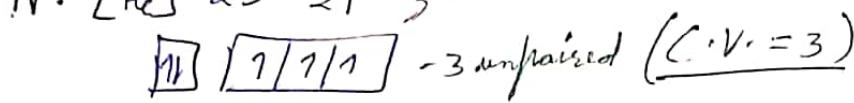


Eg ①. $CH_4, COCl_2, CCl_4$

$\rightarrow C, Si, Ge, Sn, Pb$ prefers to form 4 bonds ($C.V = 4$)

Nitrogen Family

$N: [He] 2s^2 2p^3$, cannot exist as 2d do not exist.



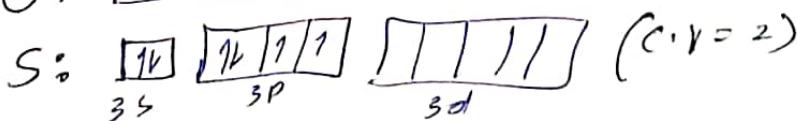
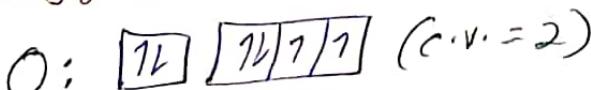
\rightarrow Nitrogen forms only ~~oxides~~

\rightarrow Nitrogen has valency 3.

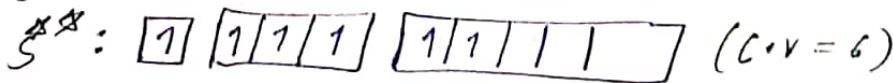
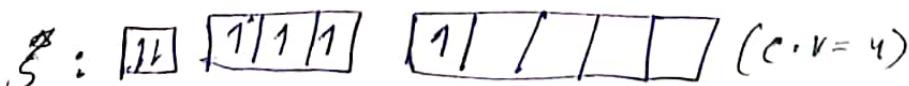
$\rightarrow P, As, Sb$ has valency both 3 & 5.

Eg ① PCl_3, PCl_5

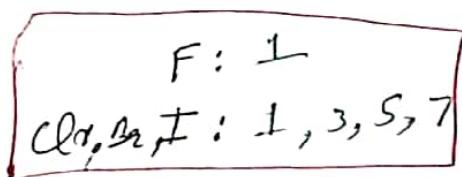
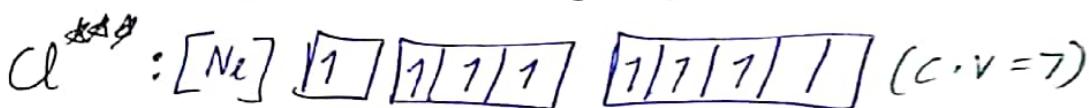
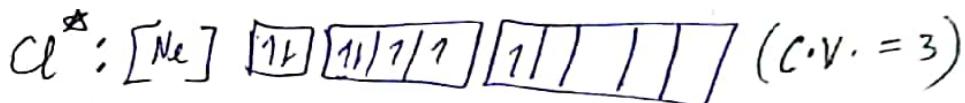
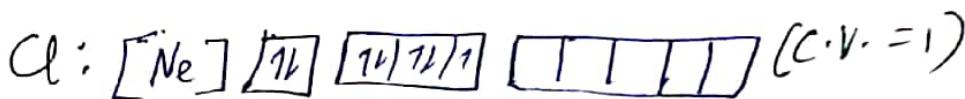
Oxygen family



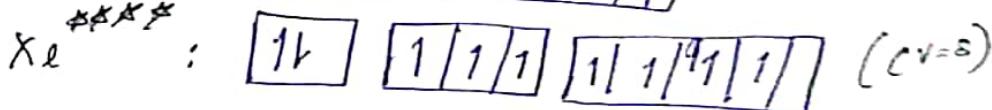
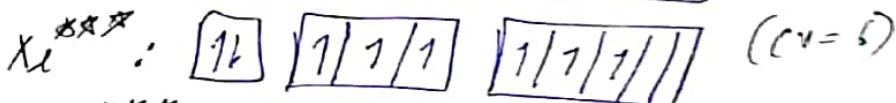
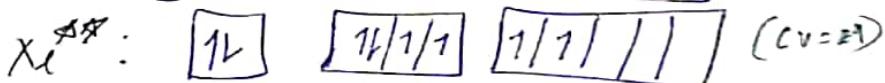
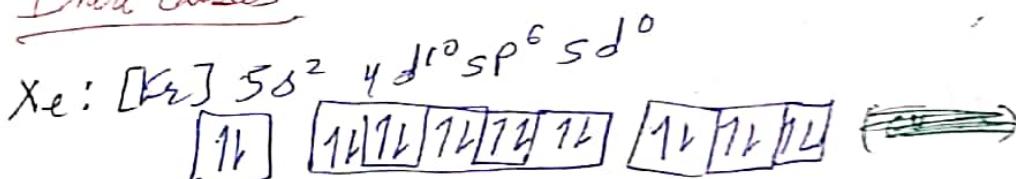
$O: 2$
 $S, Se, Te: 2, 4, 6$



Molagen Family

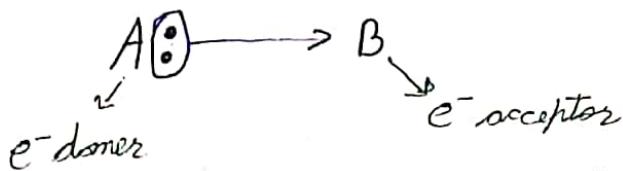


Inert Gases



3. Co-ordinate Bond (Dative Bond)

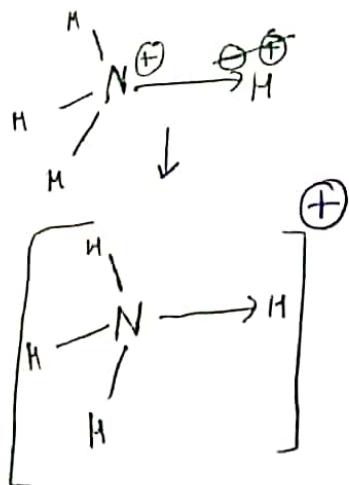
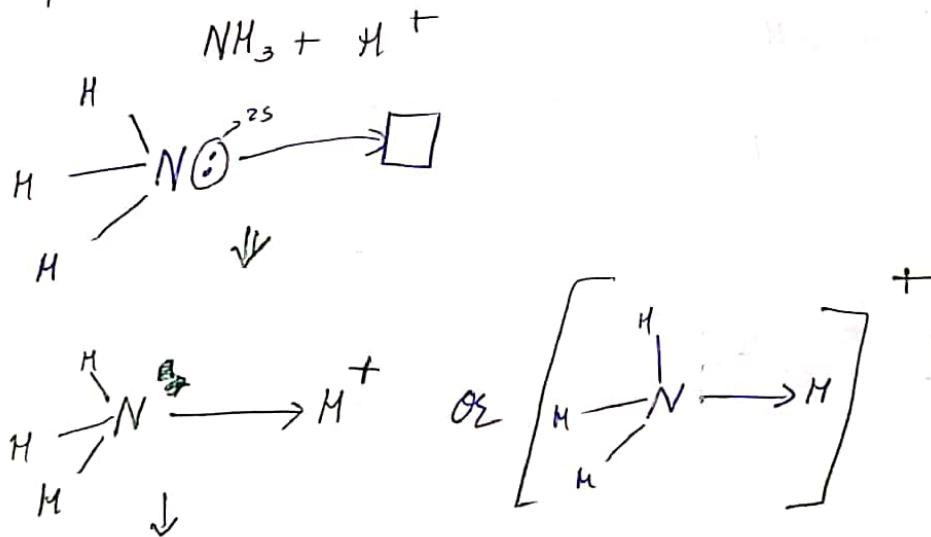
- It is formed by unequal sharing of electrons.
- It means out of two atoms between which a bond is formed, only one atom provides both the electrons that are shared between them.
- It is a type of covalent bond.



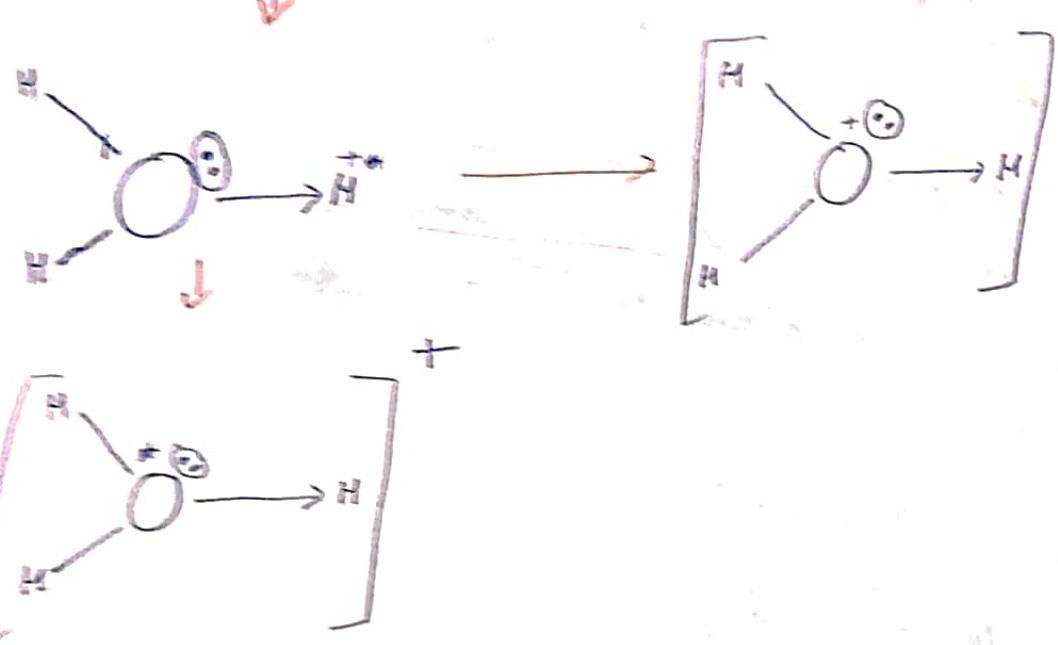
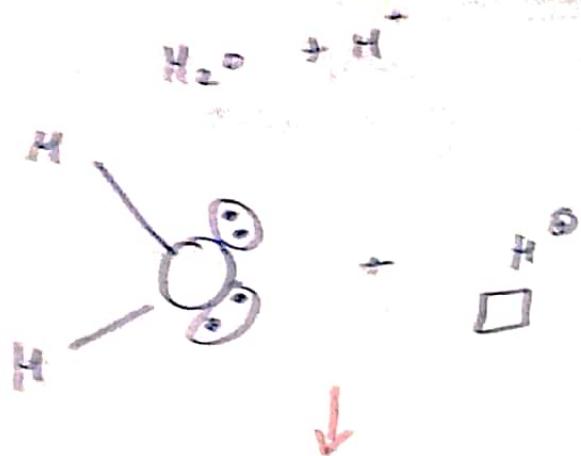
Condition to form co-ordinate bond :-

- ① One atom have at least 1 lone pair.
- ② One atom have atleast one vacant orbital.

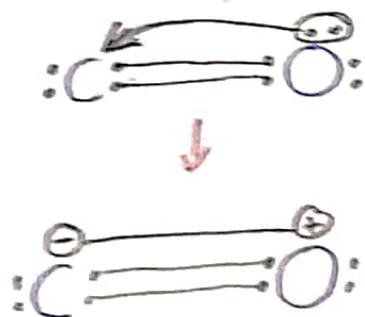
Eg NH_4^+ (Ammonium ion)



$\text{Eq} \oplus (\text{H}_3\text{O}^+)$



$\text{Eq} \oplus \text{CO}$



Lewis Dot Structure (L-D-S) :-

- According to Lewis Theory, every atom tries to obtain inert gas configuration for achieving stability.
- He proposed a theory that a chemical bond will be formed by an atom so that it has an octet configuration (except hydrogen - Duplet)

Eg ① HF



Rules to draw LDS

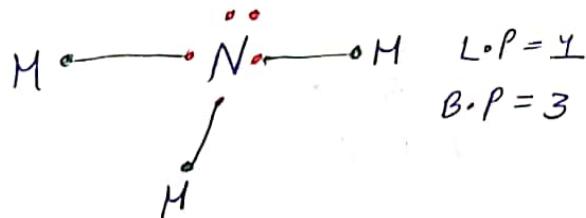
1. ~~①~~ Close ~~center~~ Central atom -

- Hydrogen & Fluorine can never be central atom
- Usually central atom is least electron negative atom and also lesser in number

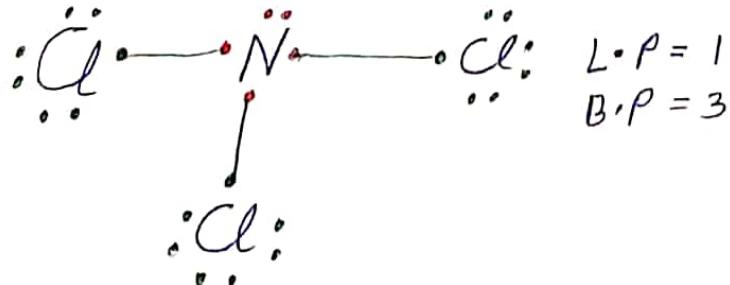
① CO_2



② NH_3



③ NCl_3



④ HCN (Cyanide)

⑤ PH₃

⑥ HNC (Iso-Cyanide)

⑦ H₂S

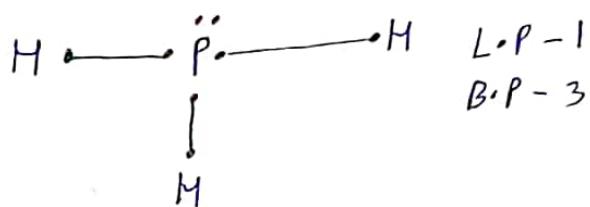
⑧ SO₂

⑨ SO₃

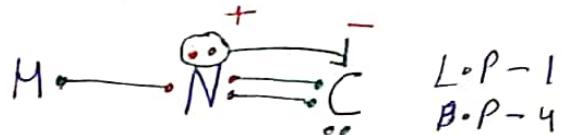
⑩



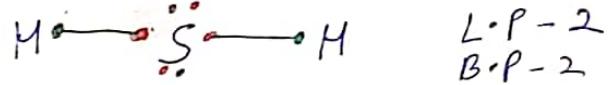
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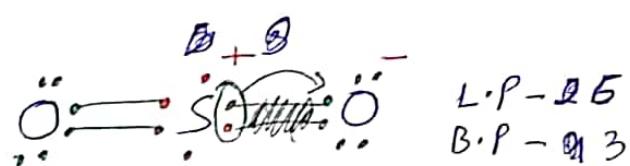
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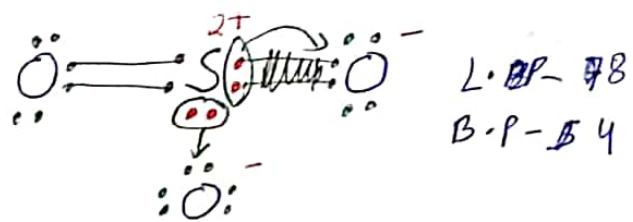
⑦



⑧



⑨



O:

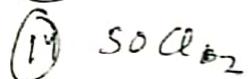
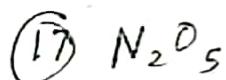
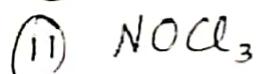
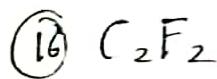
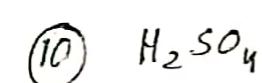
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O:

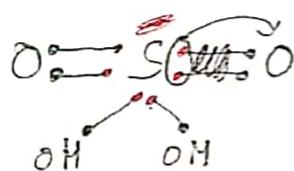
11	11	11	11
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 → To create vacant orbital for co-ordinate bond.

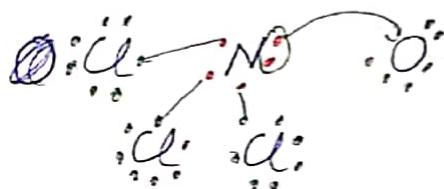
Note - ① Oxygen atom in ground state cannot form coordinate bond because it has no vacant orbital.
 ② However if 1 e⁻ of 2p subshell gets paired in another orbital of 2p subshell then a vacant orbital is created.



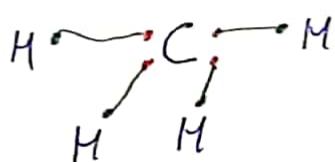
⑩



⑪



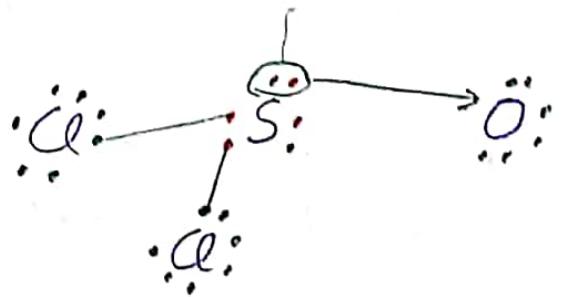
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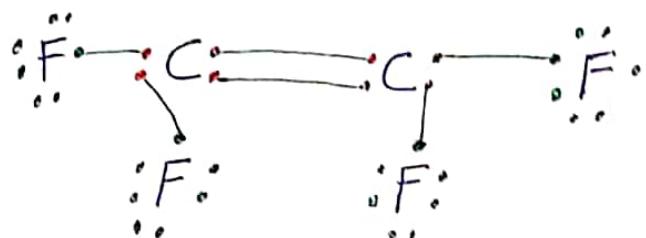
⑬



(14)



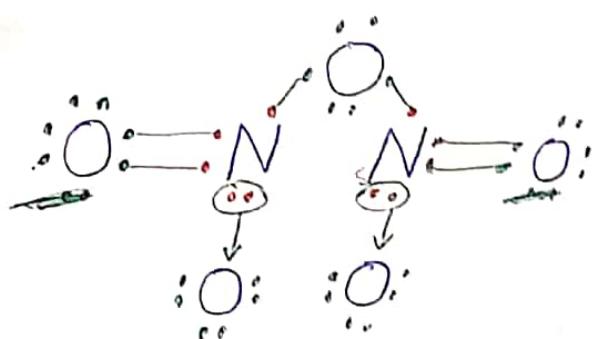
(15)



(16)



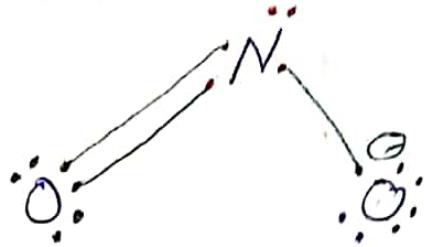
(17)

LDS of charged species

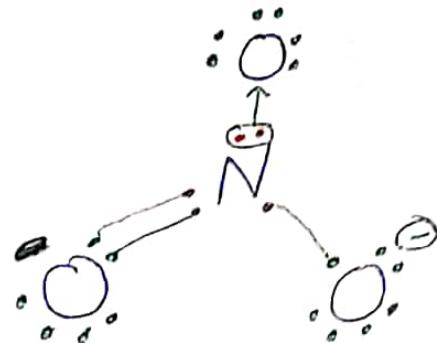
- \ominus ve charge carried by more EN atom
- \oplus ve charge carried by less EN atom.

(190)

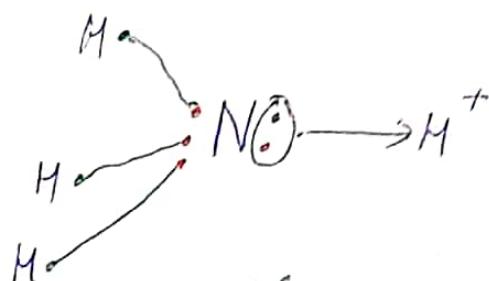
Eg ①. NO_2^-



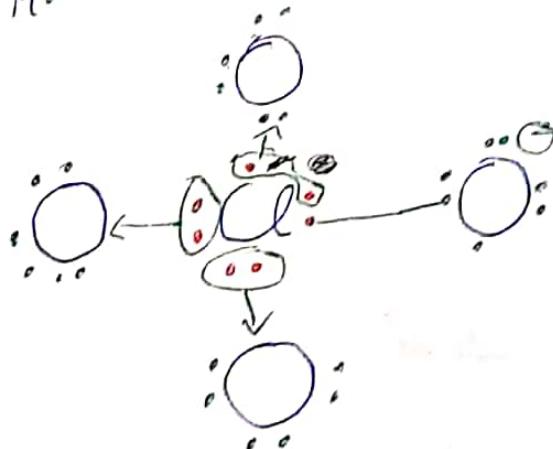
② NO_3^-

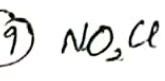
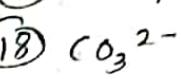
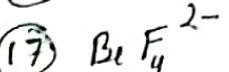
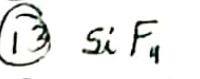
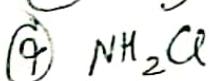
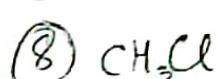
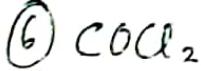


③. NH_4^+



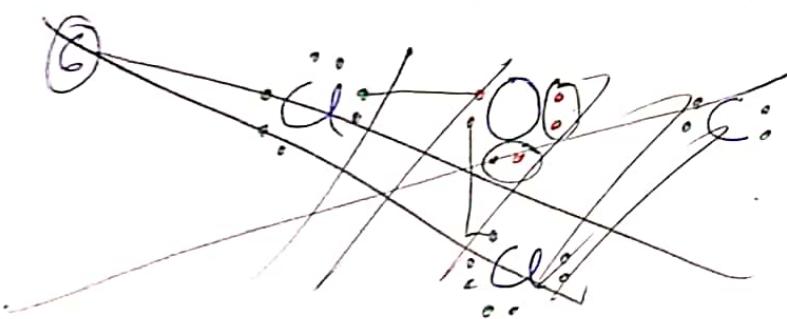
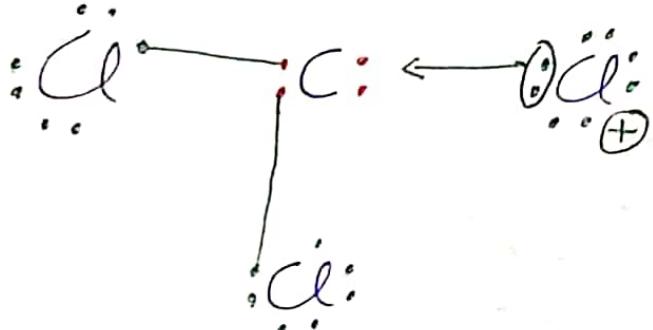
④ ClO_4^-



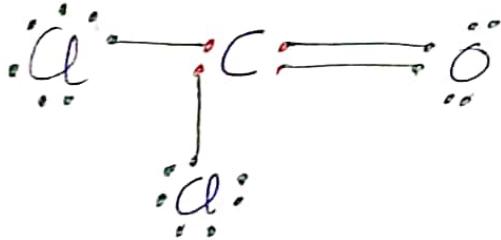


(Tildenreagent)

(5)



(6)



(7)

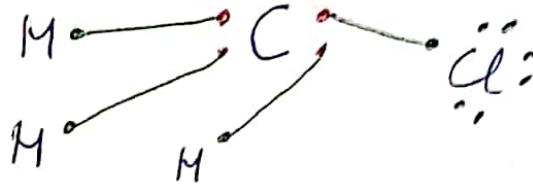


or



(192)

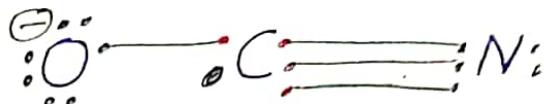
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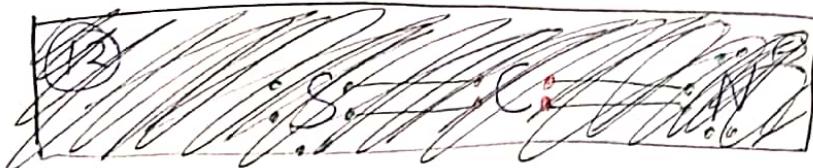
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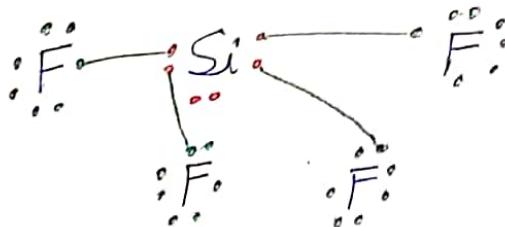
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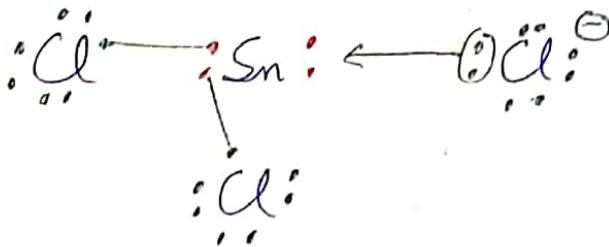
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(12)

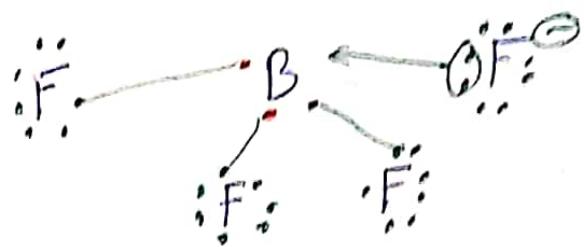


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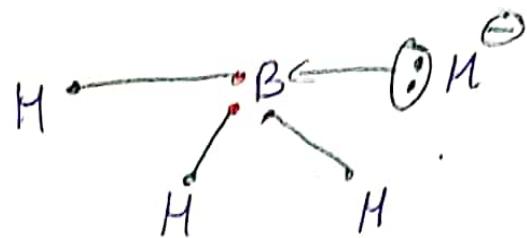


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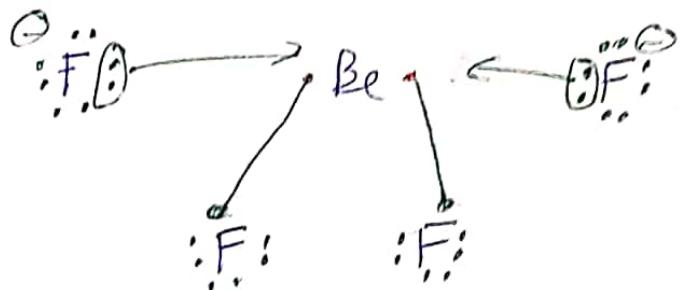
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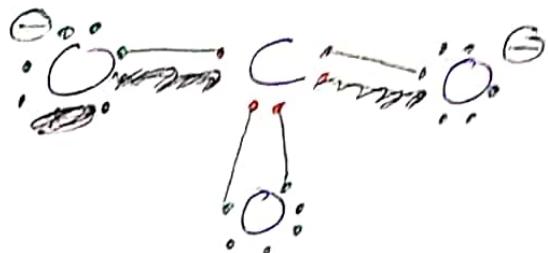
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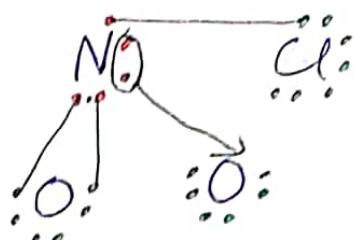
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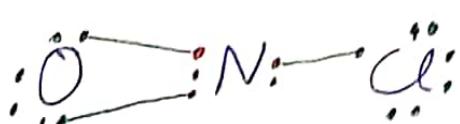
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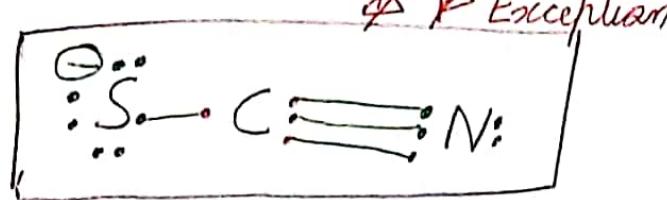


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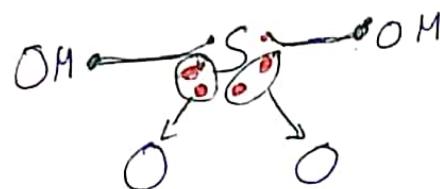
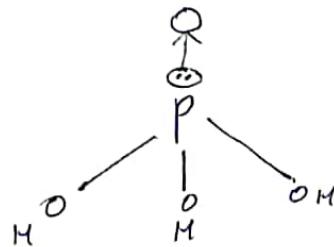


(194)

(12)

Note :-

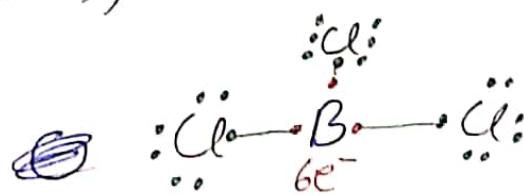
- ① If oxygen and hydrogen both are present together in a molecule. Then connect them with central atom in the form of OH bond. (few exceptions are there will be covered in oxy acids).

eg ①. H_2SO_4 eg ②. M_3PO_4 Drawbacks of Lewis Dot Structure

- ① Some compounds were discovered which were not following octet.

→ There are 3 categories

- ① hypervalent - Species in which central atom occupy less than 8 valent electrons after bonding.

eg. BF_3 , BCl_3 , BBr_3 , BT_3 , AlCl_3 

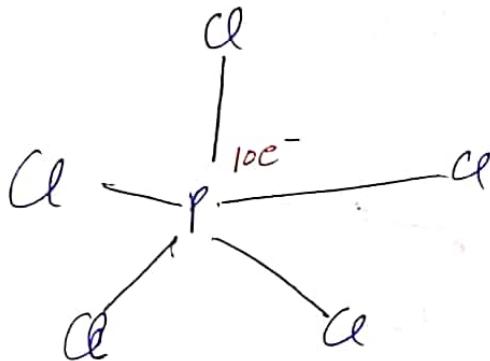
(195)

→ AlF_3 is not hypervalent because it has complete octet due to ionic nature.

→ AlF_3 is ionic but AlCl_3 , AlBr_3 & AlI_3 are covalent compound. ($\Delta E N$)

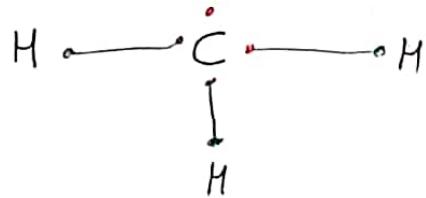
② Higher Valency - Central atom occupies more than 8 valence e⁻ after bonding.

e.g. PCl_5 , XeE_6 , XeF_4 , XeF_6 , IF_7 , IF_5



③ Odd e⁻ species - In which odd no. of ^{valence} e⁻ are present after bonding.

e.g. NO , CH , ClO_2 , NO_2 , CF_3 , ClO_3



② Lewis theory cannot explain shape of or geometry of molecule.

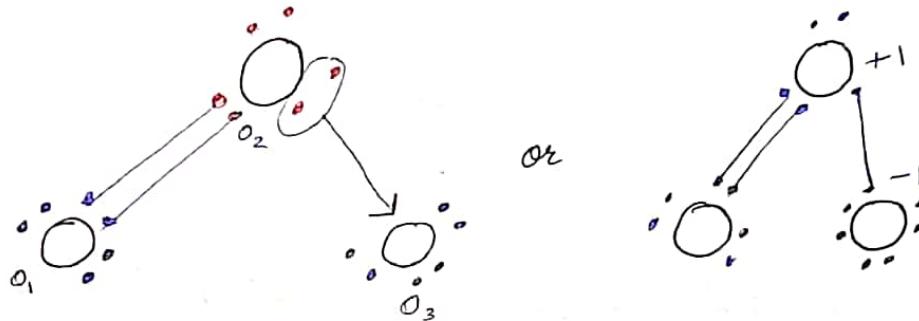
Formal Charge :-

Ch-3 chemical bonding

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

$$\{F.C.\} = \{\text{no. of valence } e^- \} - \{\text{no. of bonds formed by that atom}\} + \{-\{\text{total no. of } e^- \text{ of all P}\}\}$$

Eg ①

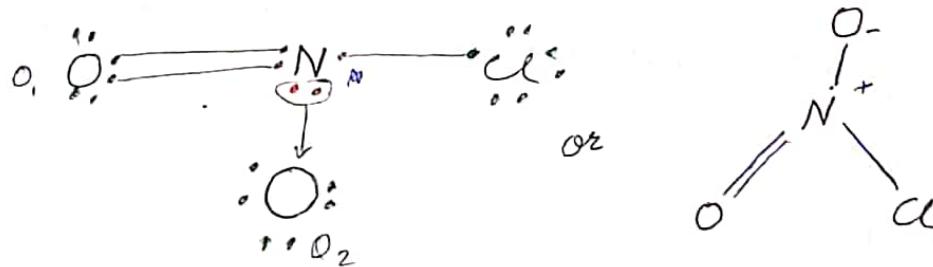


$$F.C(O_1) = 6 - 2 - 4 = 0$$

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

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

②. NO_2Cl



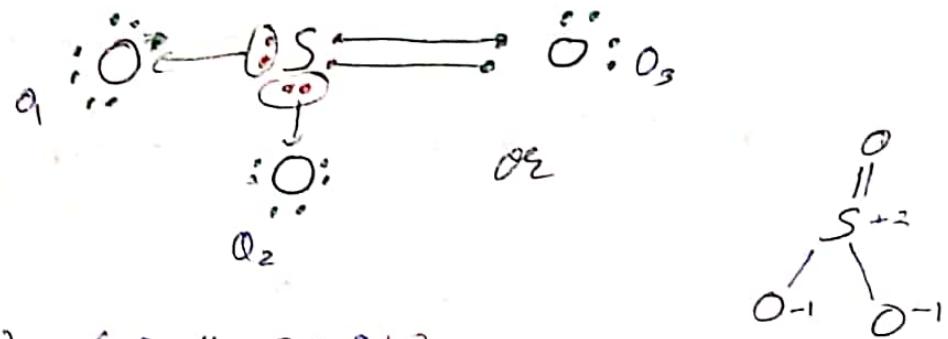
$$F.C(O_1) = 6 + -2 - 4 = 0$$

$$F.C.(O_2) = 6 - 1 - 6 = +1 - 1$$

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

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

③ SO_3



$$\text{F.C.}(\text{S}) = 6 - 4 - 0 = +2$$

$$\text{F.C.}(\text{O}_1) = 6 - 1 - 6 = -1$$

$$\text{F.C.}(\text{O}_2) = 6 - 1 - 6 = -1$$

$$\text{F.C.}(\text{O}_3) = 6 - 2 - 4 = 0$$

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

① Cl_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⁻

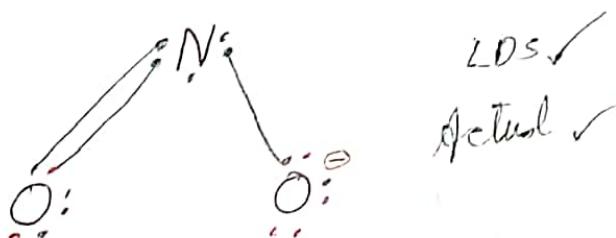
③ BCl_3 → hypervalent

④ XeO_4 → hypervalent

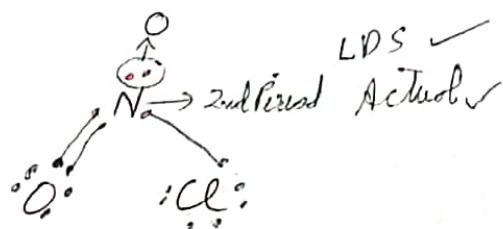
②

- 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 made 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.

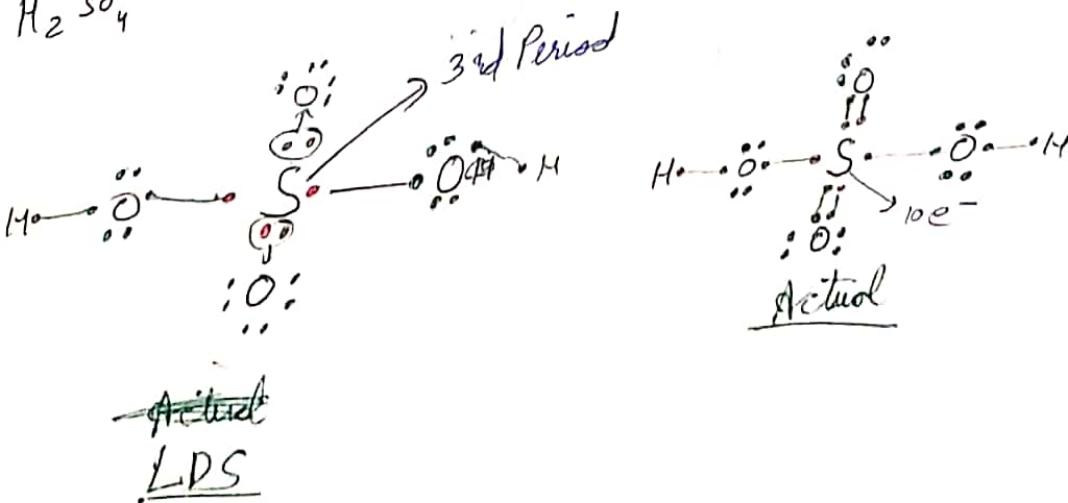
Ex ①. NO_2^-



② NO_2Cl

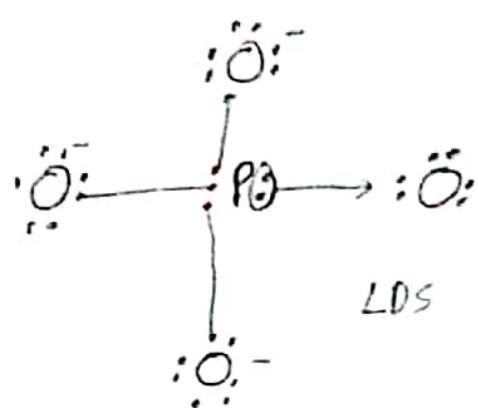
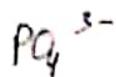


③ H_2SO_4

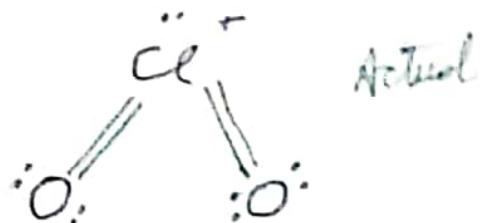
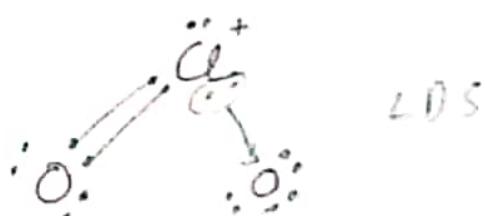


(3)

④



⑤ ClO_2^+



⑥

Q Calculate Total no. of L.P.

① SF_6

② SF_4

③ XeF_2

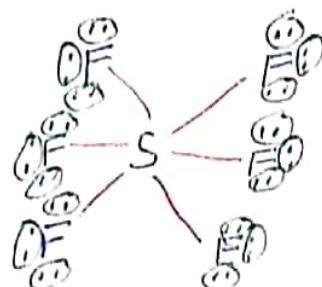
④ XeF_4

⑤ SO_2

⑥ SO_3

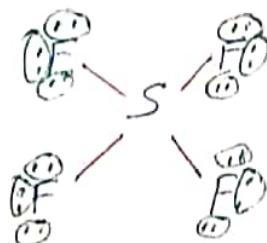
⑦ $NOCl_3$

① SF_6



$$\begin{aligned}L.P. &= 3 \times 6 \\&= 18\end{aligned}$$

② $XeSF_4$

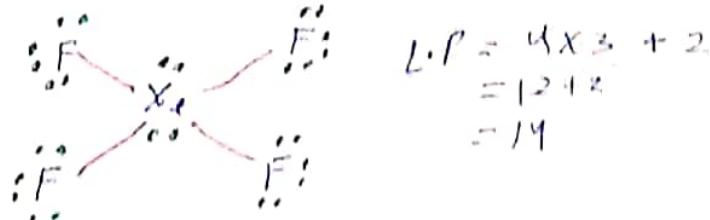


$$\begin{aligned}L.P. &= 3 + 4 + 1 \\&= 12 + 1 \\&= 13\end{aligned}$$

③ XeE_{c_2}

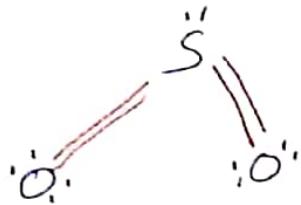


④ XeF_4



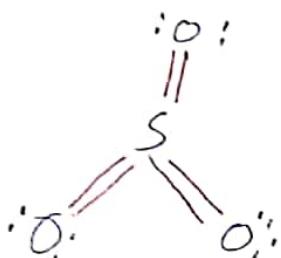
$$\begin{aligned}L.P. &= 4 \times 3 + 2 \\&= 12 + 2 \\&= 14\end{aligned}$$

⑤ SO_2



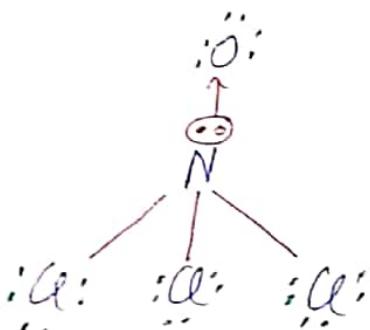
$$\begin{aligned} \text{L.P.} &= 2 \times 2 + 2 \\ &= 4 + 2 \\ &= 6 \end{aligned}$$

⑥ SO_3



$$\begin{aligned} \text{L.P.} &= 2 \times 3 \\ &= 6 \end{aligned}$$

⑦ NO_3^-



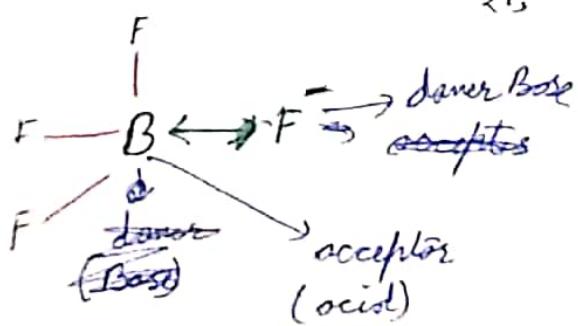
$$\begin{aligned} \text{L.P.} &= 3 + 3 \times 3 \\ &= 9 + 3 \\ &= 12 \end{aligned}$$

Note :-

① L.P. acceptor \rightarrow ~~L.P.~~ Lewis Acid

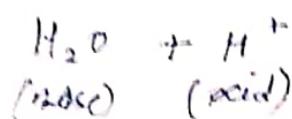
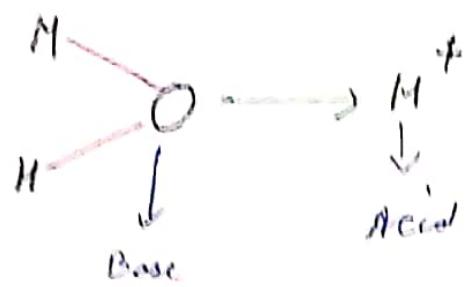
② L.P. Donor \rightarrow ~~L.P.~~ Lewis Base

③ ① BF_4^- B: $\begin{array}{c} 11 \\ 1s^2 \end{array} \quad \begin{array}{c} 1 \\ 2s^2 \end{array} \quad \begin{array}{c} 11 \\ 2p_3 \end{array}$ \rightarrow vacant orbital



④

H_2O



Oxidation Number

Oxidation Number -
→ It is the hypothetical charge on atom when each bond is assumed to be ionic.

→ Rules @for calculating oxidation number.

→ Rules @for calculating oxidation number.

① Oxidation no. of O_2
eg. $\text{Na}, \text{Cl}_2, \text{N}_2, \text{O}_2, \text{H}_2, \text{Ar}$

e.g. Na, Cl₂, H₂, O₂, etc.
Kali metals in compound state.

② Oxidation no. of ~~some~~
 $(Li, N, K, Rb, Cs, Fr) \Rightarrow -1 + 1$

e.g. LiCl, NaCl, Na_2SO_4

③ oxidation no. of Alkaline & earth metals in compound state -

$$(Bi, Mg, Ca, Sr, Ba) \rightarrow +2e$$

e.g. Ca(OH)_2 , MgCO_3 , BeCO_3 , Sr(OH)_2 , CaO

e.g. Ca(OH)_2 , Mg(OH)_2 , Be(OH)_2 , Sr(OH)_2

(4) Oxidation no. of H_2
is fixed numbered with

\rightarrow If back bonded with metals \rightarrow -1

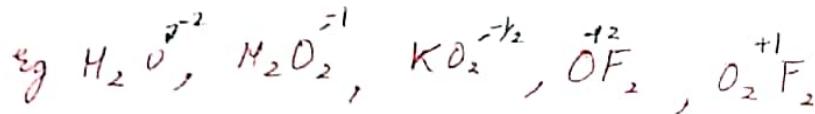
→ If bonded with other element / non-metals → +ve

e.g. CH_4 , LiH , NaNO_3 , CaF_2 , MgSO_4

⑤ Oxidation no. of oxygen atom in compound. $\Rightarrow -2$ (In general)

In peroxide (O_2^{2-}) $\Rightarrow -1$

In super oxide (O_2^{-1}) $\Rightarrow -\frac{1}{2}$



⑥ Oxidation Number of Fluorine in compound state $\Rightarrow -1$

⑦ Oxidation Number of Chlorine, Bromine & Iodine as a surrounding atom $\Rightarrow -1$

as central atom. It can increase.

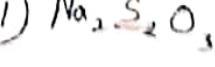
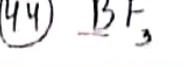
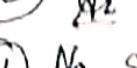
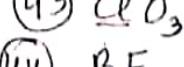
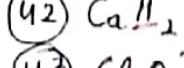
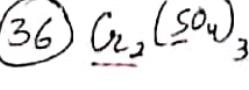
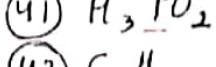
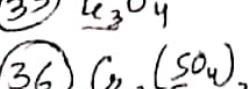
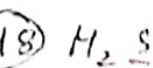
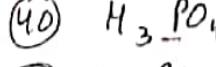
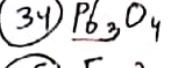
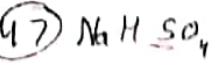
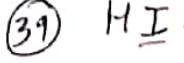
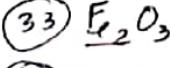
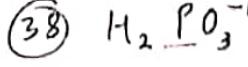
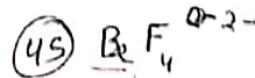
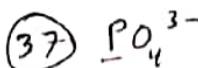
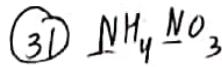
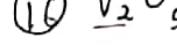
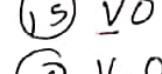
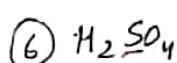
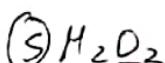
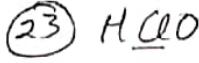
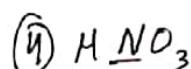
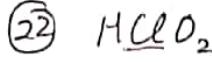
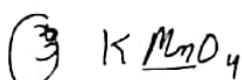
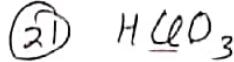
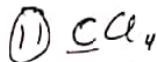
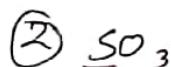
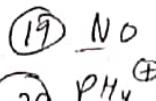
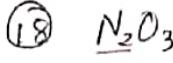
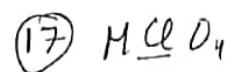
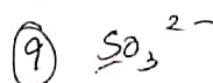
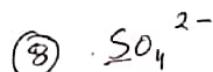
H.W.



$$S + 2(-2)$$

$$S = -4$$

$$\boxed{S = +4} \checkmark$$



㉟

① 50,

$$S + 40 = 50$$

$$S + 40 - 40 = 50 - 40$$

$$\boxed{S = 10} \checkmark$$

② $M_1 = 8 - 1$

$$\boxed{M_1 = 7} \checkmark$$

③ $2P - 8 + 3 = 0$

$$\begin{aligned} 2P - 8 + 3 &= 0 \\ S &= 8 - 2 \\ \boxed{S = 6} &\checkmark \end{aligned}$$

④ $P + 1 - 6 + N = 0$

$$N = 6 - 1$$

$$\boxed{N = 5} \checkmark$$

⑤ $2 \cdot 0 + 2 = 20 = 0$

$$\begin{aligned} 20 &= 2 \\ 20 - 20 &= 0 - 1 \\ \boxed{0 = -1} &\checkmark \end{aligned}$$

⑥ ⑦ $1 + 20 = 0$

$$20 = -1$$

$$\boxed{0 = -1} \checkmark$$

⑧ $S - 8 = -2$

$$S = 8 - 2$$

$$\boxed{S = 6} \checkmark$$

⑨ \cancel{S}

$$S - 2 = 8 - 6$$

$$\boxed{\cancel{S}} \checkmark$$

$$S = 6 - 2$$

$$\boxed{S = 4} \checkmark$$

⑩ $N + 4 = 1$

$$N + 4 - 4 = 1 - 4$$

$$\boxed{N = -3} \checkmark$$

⑪ $\boxed{P + 1} \checkmark$

⑫ $M_1 - 8 = -2$

$$\boxed{M_1 = 6} \checkmark$$

⑬ $2G_2 P - 14 = -2$

$$\begin{aligned} 2G_2 &= 12 \\ G_2 &= 6 \end{aligned} \boxed{G_2 = 6} \checkmark$$

⑭ $P + 2 = 0$

$$\boxed{P = -2} \checkmark$$

⑮ $V - 2 = 0$

$$\boxed{V = 2} \checkmark$$

⑯ $2V - 10 = 0$

$$\begin{aligned} V &= 10 \\ V &= 5 \end{aligned} \boxed{V = 5} \checkmark$$

⑰ $1 + Cl - 8 = 0$

$$\boxed{Cl = 7} \checkmark$$

⑱ $2N - 6 = 0$

$$\boxed{N = 3} \checkmark$$

⑲ $N - 2 = 0$

$$\boxed{N = 2} \checkmark$$

⑳ $P + 4 = 1$

$$\boxed{P = -3} \checkmark$$

㉑ $Cl + G = 0$

$$\boxed{Cl = G} \checkmark$$

㉒ $1 + Cl - 4 = 0$

$$\boxed{Cl = 3} \checkmark$$

㉓ $Cl = 1$

Note -

oxidation State

① For any element range of $\overset{1}{\underset{-1}{\text{O.S}}}$ \rightarrow N to N-8 (both included)
 \rightarrow no of valence e⁻

Eg ① N: N=5

$$N: \{ +5 \text{ to } -3 \}$$

② F; S: N=c

$$S: \{ +6 \text{ to } -2 \}$$

③ O: N=c

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

$$O: \{ -2 \text{ to } +2 \} \text{ (Actual) } \cancel{\text{Exception}}$$

④ Cl: N=7

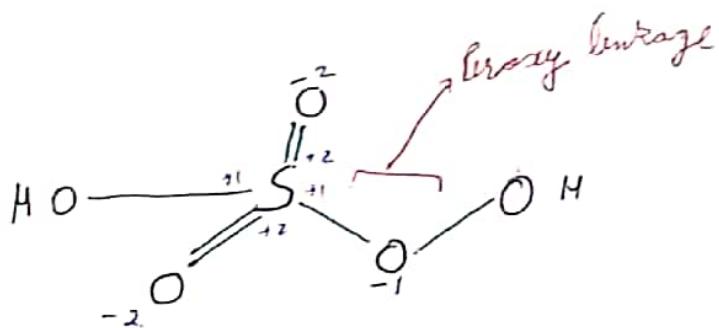
$$Cl: \{ -1 \text{ to } +7 \}$$

Eg H_2SO_4 (con's acid)

$$2+x-10=0$$

$$\cancel{x=10-2}$$

$x=8$ \rightarrow but had only 7! ($-2 \neq 0$)
 So wrong

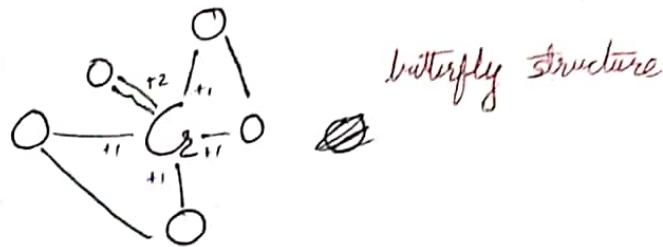


$$\begin{array}{|c|} \hline 2+2+1+1=6 \\ \hline \text{O.S of } S=6 \\ \hline \end{array}$$

(10)

Eg. CrO_5

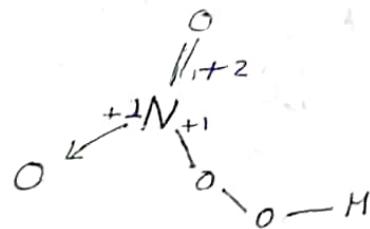
$$x - 10 = 0 \\ x = 10 \quad \text{but chromium has only } 6 \text{ valence e}^-$$



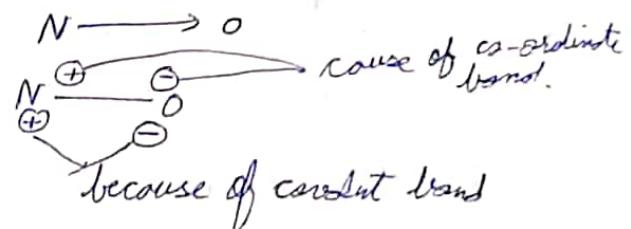
$$2+1+1+1+1 = \boxed{6}$$

Eg. HNO_4

$$N = 8 - 1 \\ = 7 \quad \times$$



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



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

- May or May not equal 0

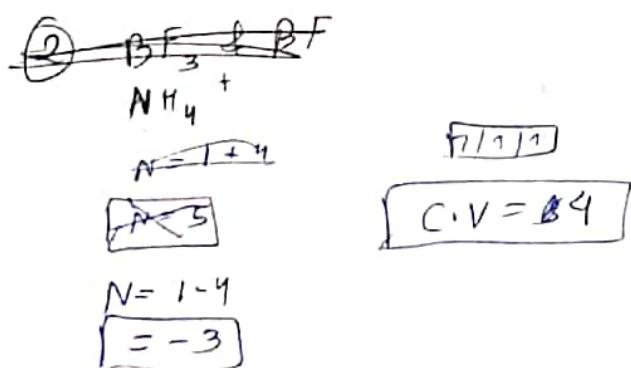
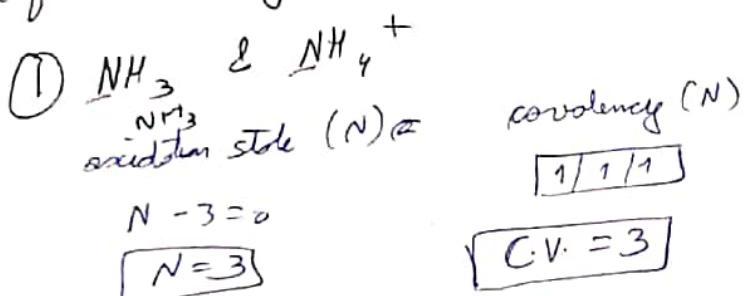
Covalency

- ① No. of unpaired e^- in ground state or excited state.
- ② Can be negative or fractional.
- ③ Depends on the no. of bonds

Oxidation state

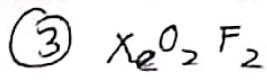
- ① Permanent formal charge.
- ② Can be Θ ve or fractional.
- ③ Does not depend on no. of bonds.

Q find covalency and oxidation state of



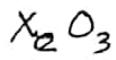
$$O.S = 3 \quad O.S = 3$$

$$C.V = 3 \quad C.V = 4$$



$$O.S = 6$$

$$C.V = 6$$



$$O.S = 6$$

$$C.V = 6$$

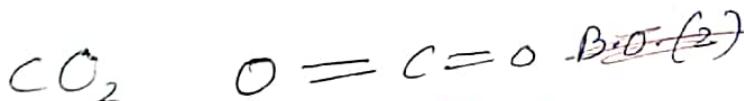
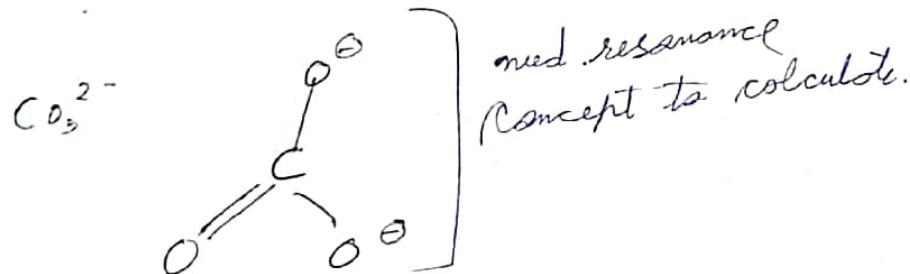
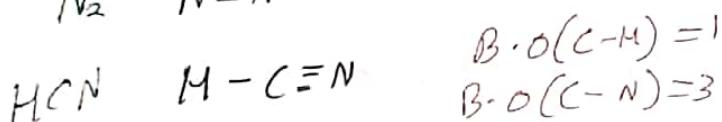
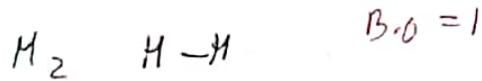


$$O.S = 8$$

$$C.V = 8$$

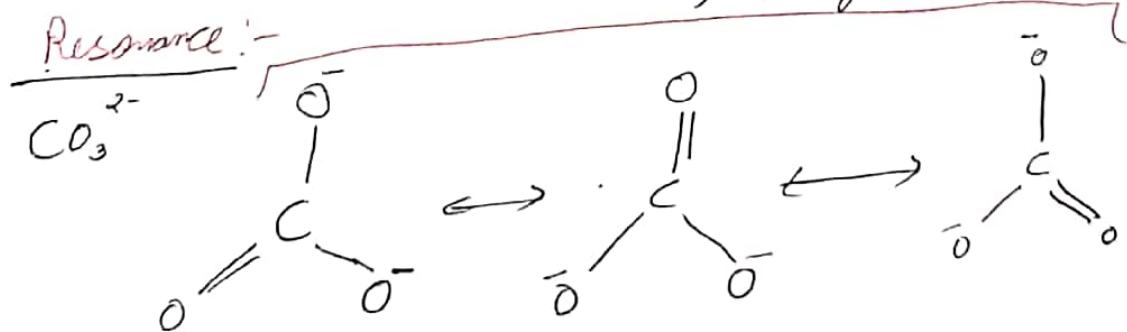
Bond Order (B.O).:-

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

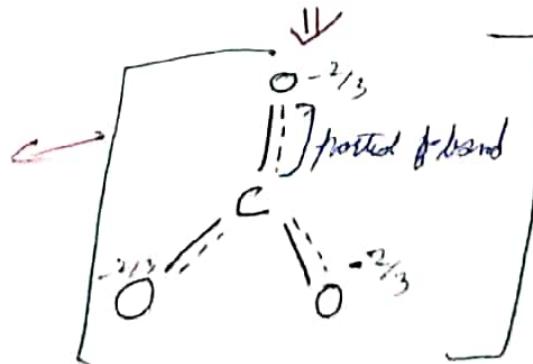


$$B.O.(\cancel{2}) = 2$$

resonating structures or canonical



more stable



Resonance Hybrid.

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

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

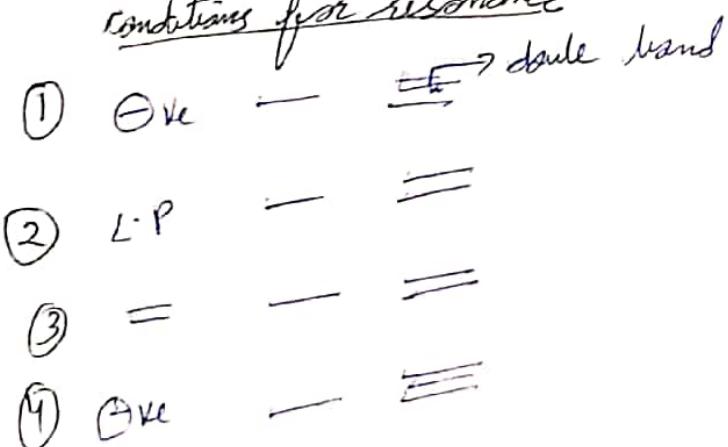
→ σ is the first bond formed between two atoms
 → π is the second bond formed between two atoms.

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 'Resonance Hybrid of the Structure'

- This phenomenon is called resonance and depends contributing structures is called resonating or canonical structures.
- Resonance stabilize the molecule as the energy of Resonance hybrid is less than energy of any single canonical structure.

conditions for resonance

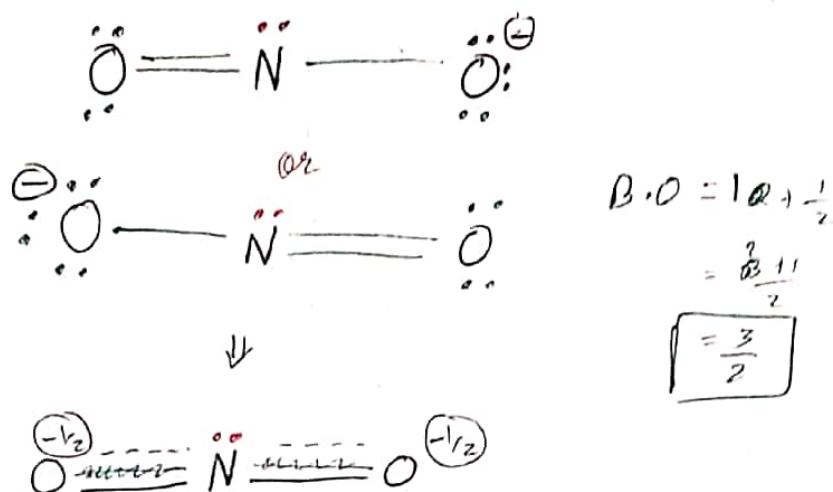


- 24 $\boxed{c_2 = 6}$ ✓
 25 $\boxed{1}$ +3
 26 $\boxed{-3}$ ✓
 27 $\boxed{-3}$ ✓
 28 $\boxed{3}$ ✓
 29 $\boxed{5}$ ✓
 30 $\boxed{16}$ ✓
 31 $\boxed{1}$ -3, +5
 32 $\boxed{2}$ ✓
 33 $\boxed{3}$ ✓
 34 $\boxed{\frac{8}{3}}$ ✓
 35 $\boxed{\frac{+8}{3}}$ ✓
 36 $\boxed{1, 6}$ +3, 6
 37 $\boxed{5}$ ✓
 38 $\boxed{3}$ ✓
 39 $\boxed{7}$ ✓
 40 $\boxed{5}$ ✓
 41 $\boxed{1}$ ✓
 42 $\boxed{-1}$ ✓
 43 $\boxed{5}$ ✓

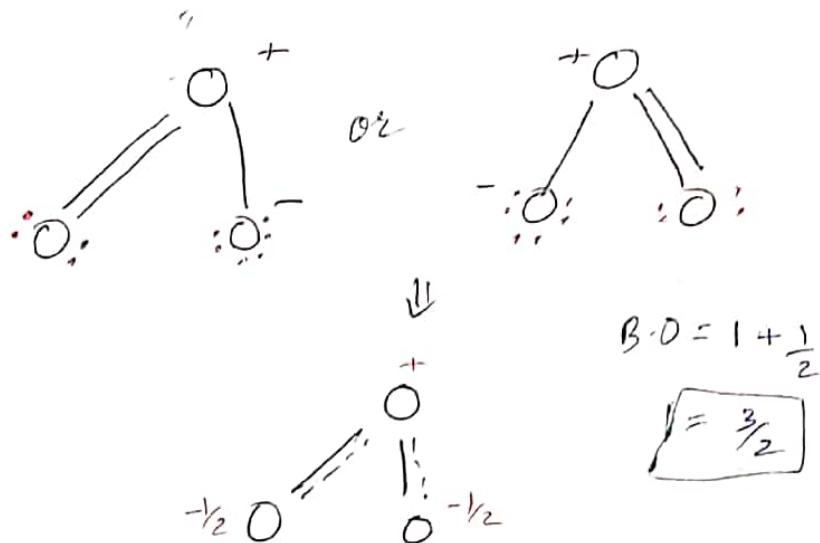
- 44 $\boxed{3}$ ✗✓
 45 $\boxed{2}$ ✓
 46 $\boxed{2}$ ✓
 47 $\boxed{6}$ ✓
 48 $\boxed{-2}$ ✓
 49 $\boxed{-\frac{1}{3}}$ ✓
 50 $\boxed{0}$ ✓
 51 $\boxed{2}$ ✓

(31)

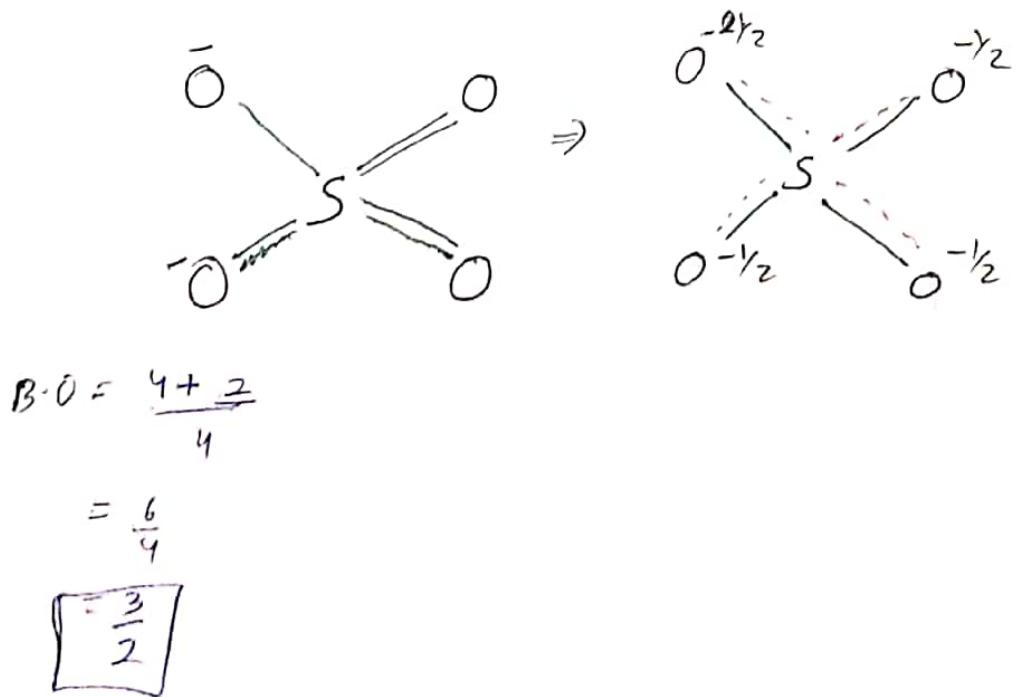
Eg(1) NO_2^-



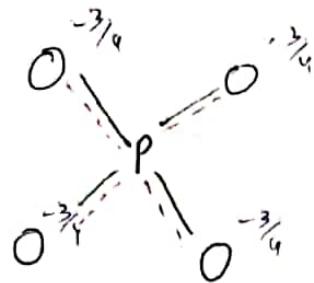
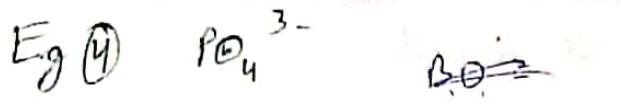
Eg(2) O_3



Eg(3) SO_4^{2-}

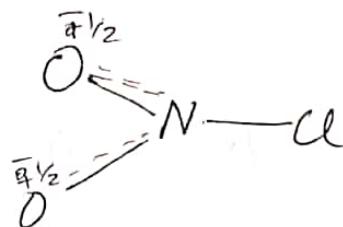
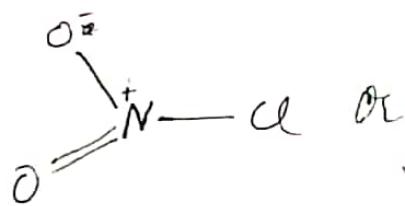
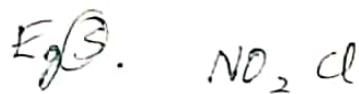


(16)



$$\text{B.O} = \frac{\cancel{8}4+1}{4}$$

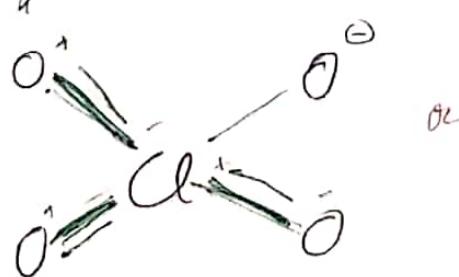
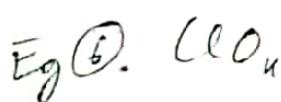
$$\boxed{= \frac{5}{4}}$$



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

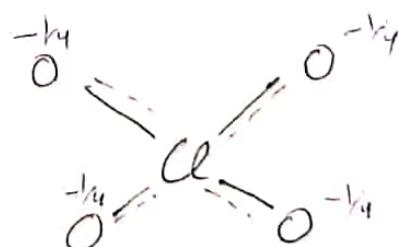
$$\text{B.O} (\text{O-N}) = \frac{2+1}{2}$$

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



$$\text{B.O} = \boxed{5} \frac{9+03}{4}$$

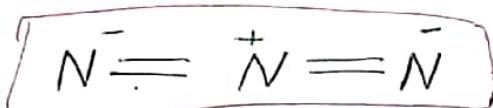
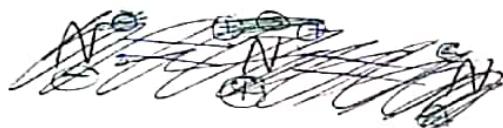
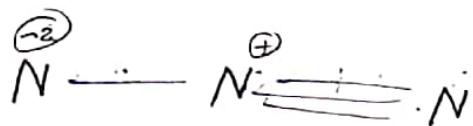
$$\boxed{= \frac{7}{4}}$$



Q Draw the Resonating Structure Of Azide ion ?
(N₃⁻)



B.D.



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



Note:-

①
$$\boxed{B.O \propto B.S \propto B.E \propto B.P.E \propto \frac{1}{B.L}}$$

B.O → Bond Order

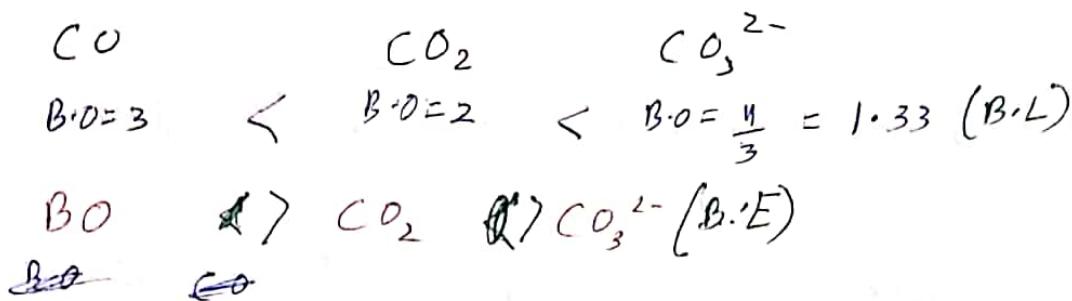
B.S → Bond Strength

B.E → Bond Energy

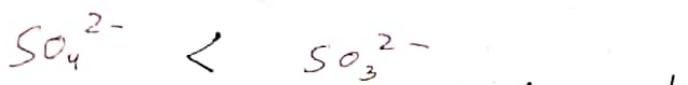
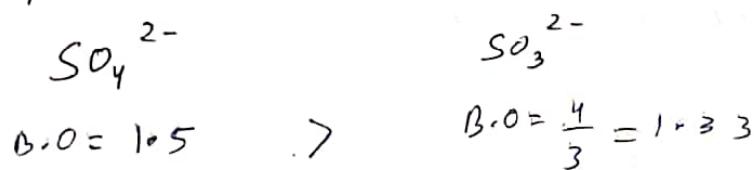
B.D.E → Bond Dissociation energy

B.L → Bond Length

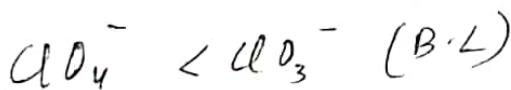
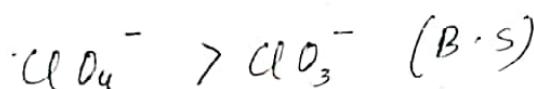
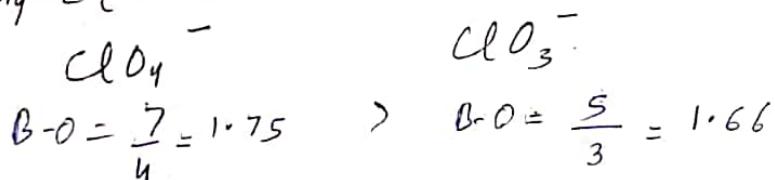
Q Compare bond length in (O-C-O)



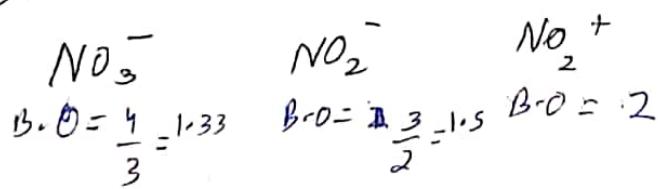
Q Compare (S-O) Bond Length



Q Compare (Cl-O) Bond strength & length

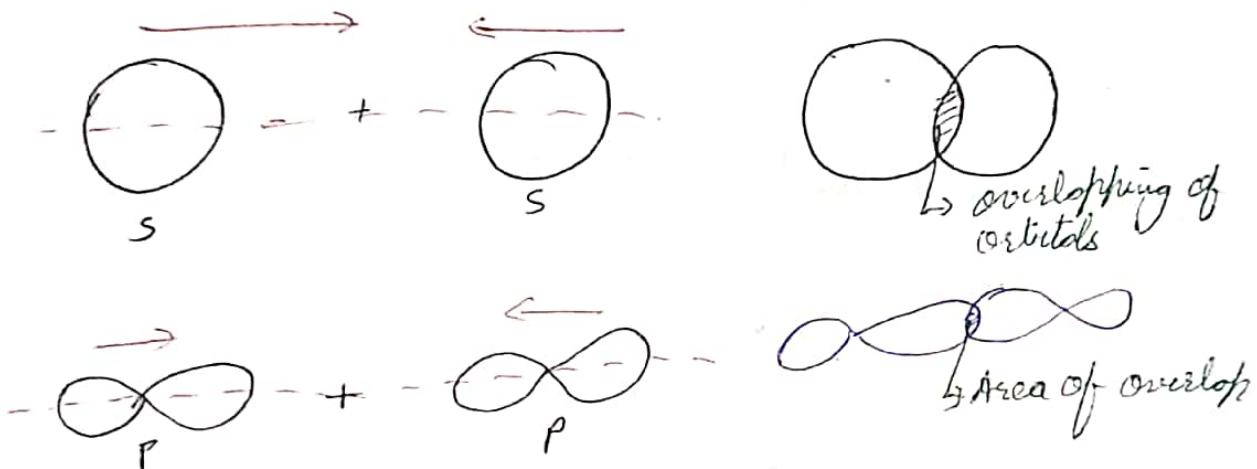


Q (N-O) bond length

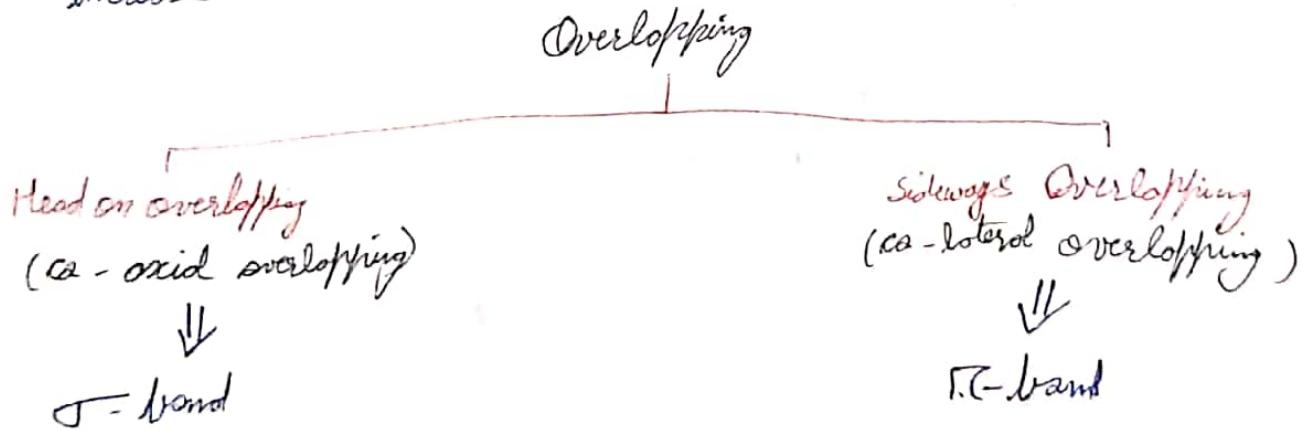


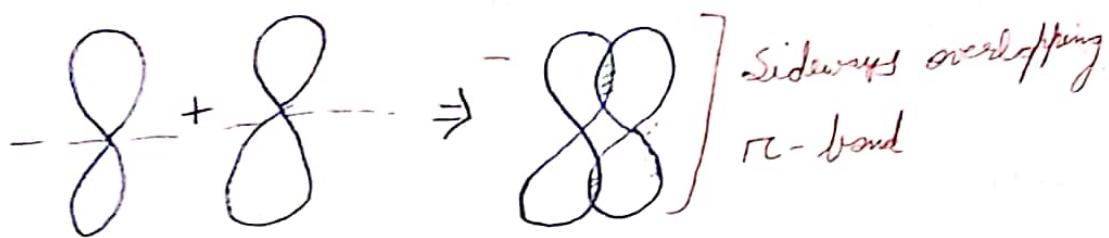
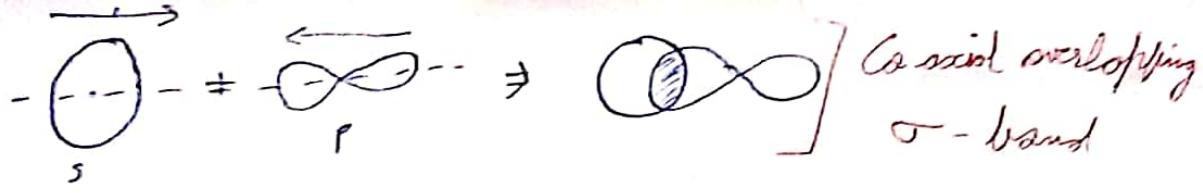
Valence Bond Theory (VBT)

- It Explains how covalent bond will be formed.
- Structures ~~were~~ based on valency.
- It is based on overlapping of atomic orbitals.
- A bond is formed by overlapping of pure atomic orbitals in a specified orientation.



Note - If overlapping area increases then covalent strength increases

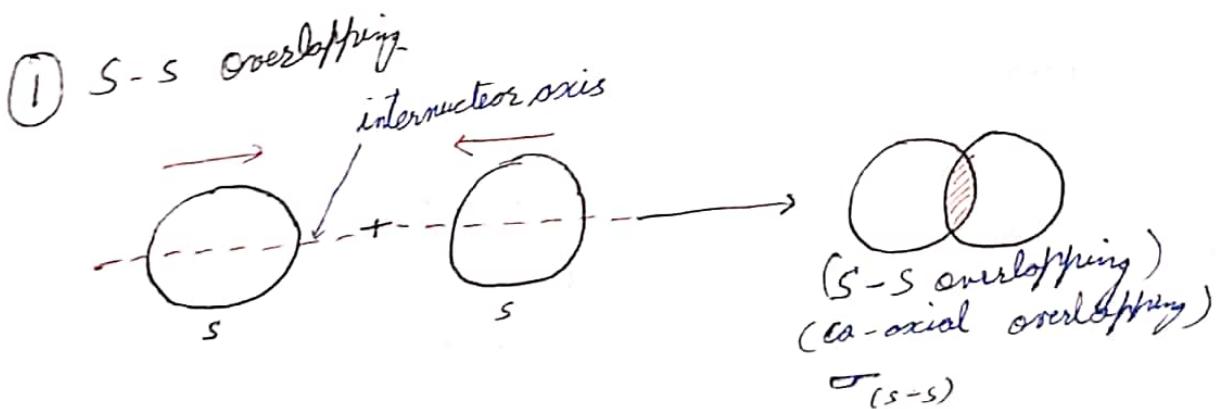




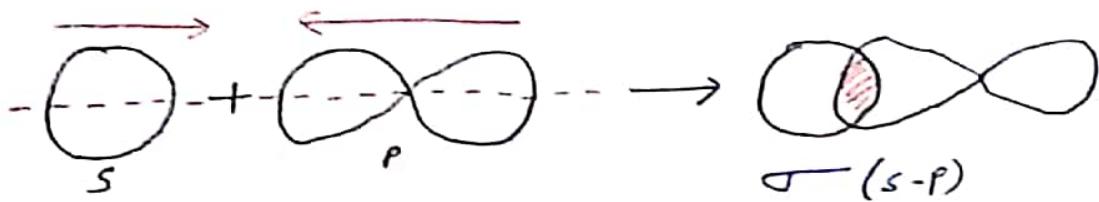
Note → σ bond is stronger than π bond as overlapping area is more, electron density is more.

Covalent Bond

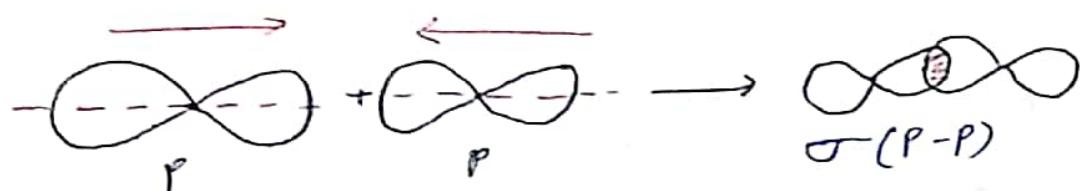
- σ bond
- ① S-S overlapping
- ② S-P overlapping
- ③ P-P overlapping
-
- π - bond
- ① P-Orbital overlapping
- ② P-d overlapping
- ③ d-d overlapping.



② $s - sp$ overlapping

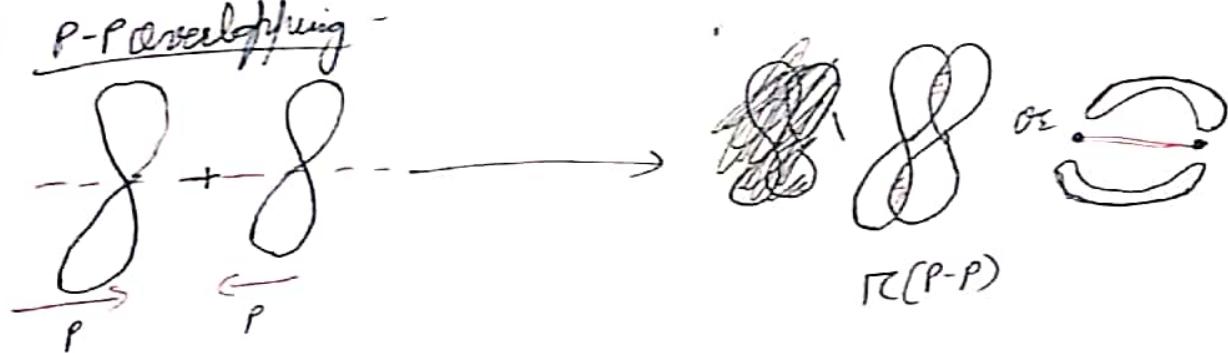


③ $p - p$ overlapping

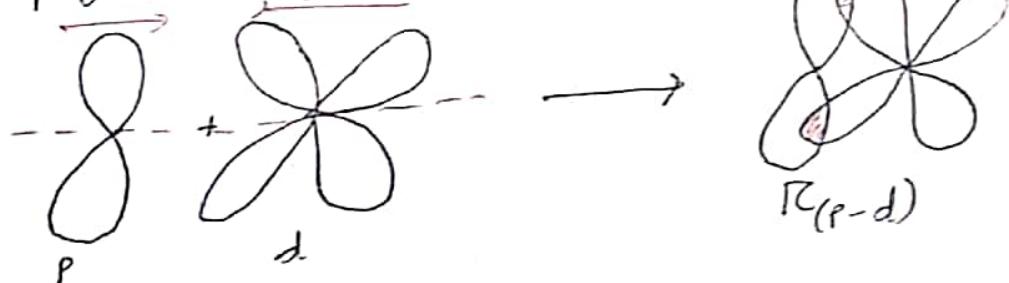


π -bonds

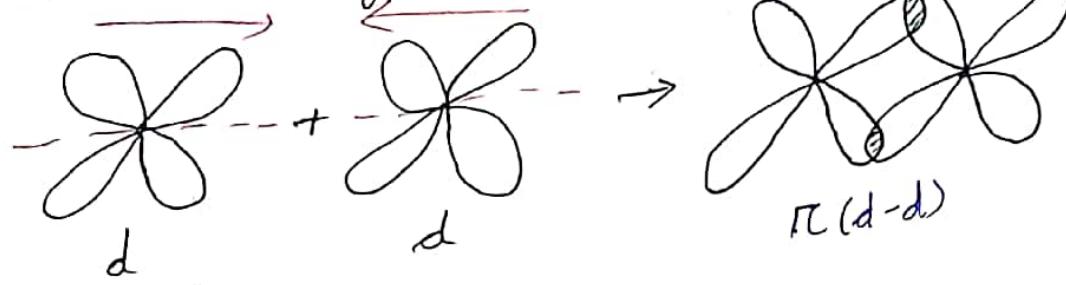
④ $p - p$ overlapping



⑤ $p - d$ overlapping



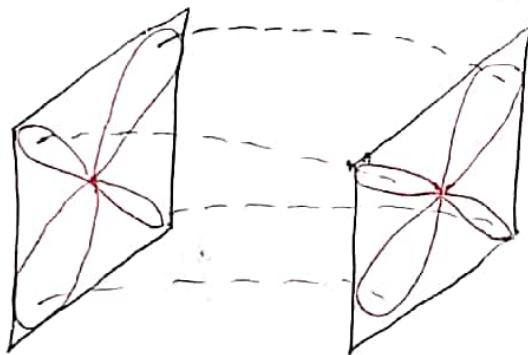
⑥ $d-d$ overlapping



Delta Bond (δ):-

(Type of π bond)

→ It is a special type of π bond in which 4 lobes are interoverlapped known as ~~delta~~ δ bond.



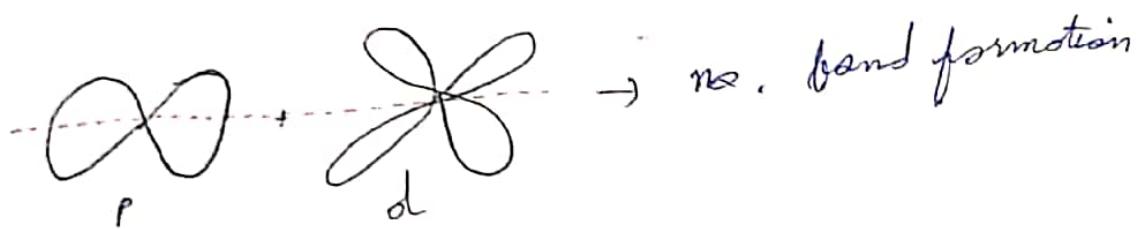
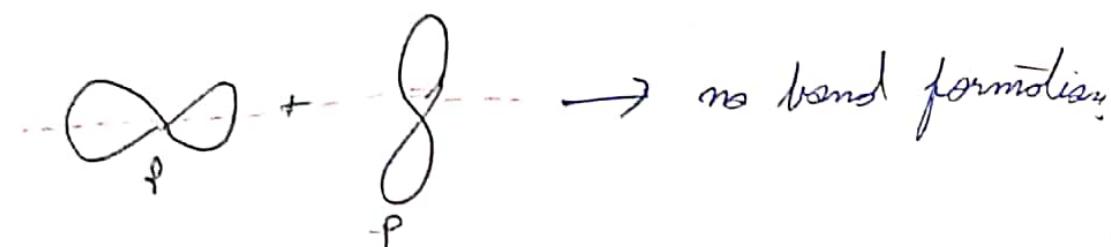
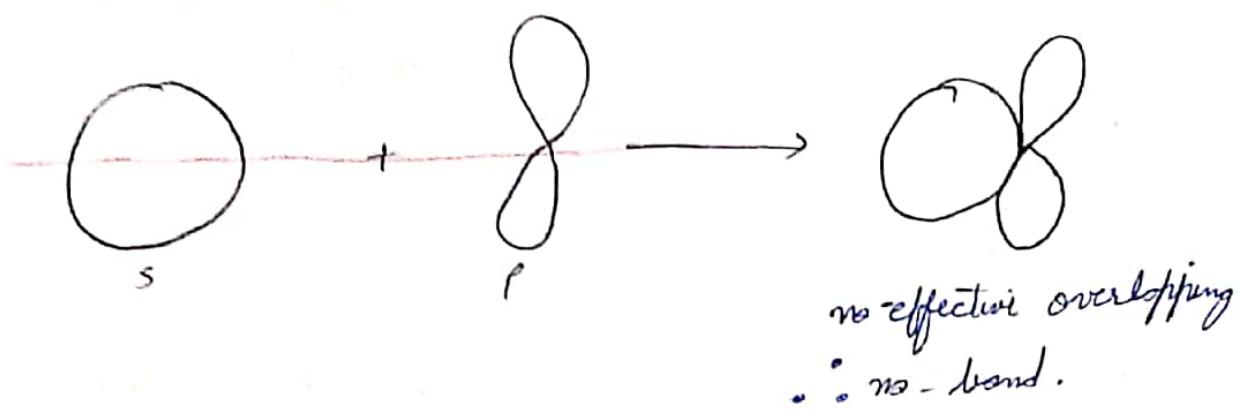
Note:- w.r.t one orbital

- ① 1. lobes Interacted
- 2. lobes Interacted
- 3. 4. lobes Interacted

Type of bond
 δ bond

π bond
 σ bond.

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



Note:-

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

Molecular orbit① $S + P_x$ Interatomic axis $x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$ Type of bond σ
 x
 x ② $S + P_y$ $x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$ x
 σ
 x ③ $S + P_z$ $x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$ x
 x
 σ ④ $P_x + P_z$ $x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$ σ
 π
 π ⑤ $P_{ay} + P_{gy}$ $x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$ π
 σ
 π ⑥ $P_{az} + P_{gz}$ $P_x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$ π
 π
 σ ⑦ $P_c + P_g$ $x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$ x
 x
 x

$$\textcircled{8} \quad p_y + p_z$$

$x \rightarrow$
 $y \rightarrow$
 $z \rightarrow$

x
 x
 x

$$\textcircled{9} \quad p_x + d_{xy}$$

$x \rightarrow$
 $y \rightarrow$
 $z \rightarrow$

x
 Γ
 x

$$\textcircled{10} \quad p_y + d_{yz}$$

$x \rightarrow$
 $y \rightarrow$
 $z \rightarrow$

x
 x
 Γ

$$\textcircled{11} \quad d_{xy} + d_{xz}$$

$x \rightarrow$
 $y \rightarrow$
 $z \rightarrow$

Γ
 Γ
 δ

$$\textcircled{12} \quad d_{yz} + d_{yx}$$

$x \rightarrow$
 $y \rightarrow$
 $z \rightarrow$

δ
 Γ
 Γ

$$\textcircled{13} \quad d_{zx} + d_{xz}$$

$x \rightarrow$
 $y \rightarrow$
 $z \rightarrow$

Γ
 δ
 Γ

$$\textcircled{14} \quad d_{zy} + d_{yz}$$

$x \rightarrow$
 $y \rightarrow$
 $z \rightarrow$

x
 x
 x

$$\textcircled{15} \quad d_{yz} + d_{zx}$$

$x \rightarrow$
 $y \rightarrow$
 $z \rightarrow$

x
 x
 x

$$\textcircled{16} \quad d_{xy} + d_x d_z$$

$x \rightarrow$
 $y \rightarrow$
 $z \rightarrow$

x
 x
 x

(26)

H.W. (29-05-2024)

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

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

S-1 $\{16\}$

S-2 \emptyset ~~$\{3, 4, 13, 18, 19\}$~~

Note -

- ① Two perpendicular orbitals can never form any bond.
- ② π bond is formed only after ~~signs~~ σ bond.
- ③

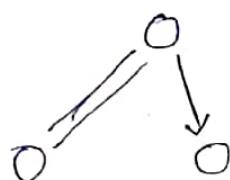
Q Calculate π bond & σ bond?

① CO_2

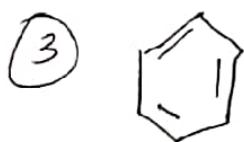
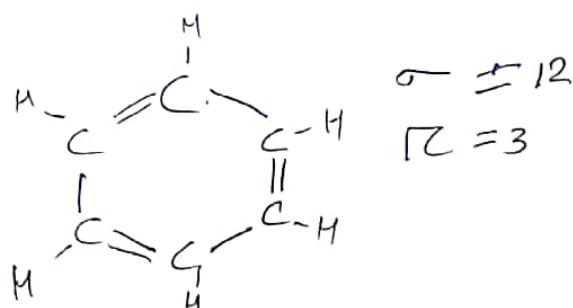


π bonds - 2
 σ bonds - 2

② O_3



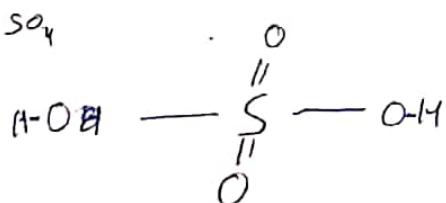
$\sigma = 2$
 $\pi = 1$



~~sigma~~
~~pi~~

④

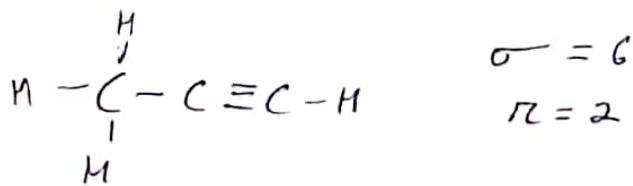
H_2SO_4



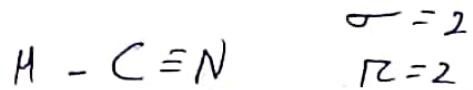
$\sigma = 16$
 $\pi = 2$

⑤

$\text{H}_3\text{C}-\text{C}\equiv\text{CH}$



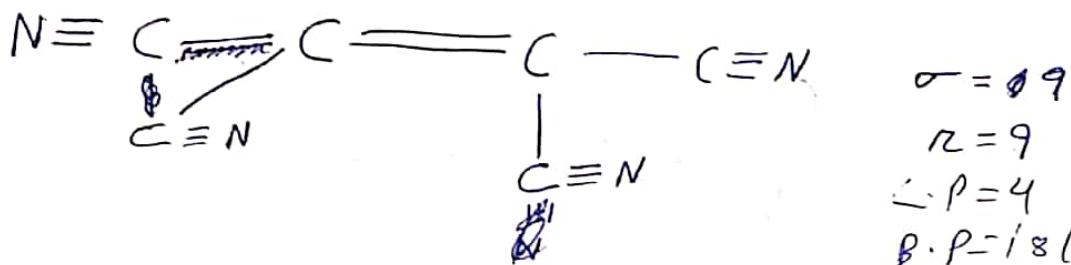
⑥ HCN



$$\sigma = 2$$

$$r = 2$$

⑦ $\text{C}_2(\text{CN})_4$



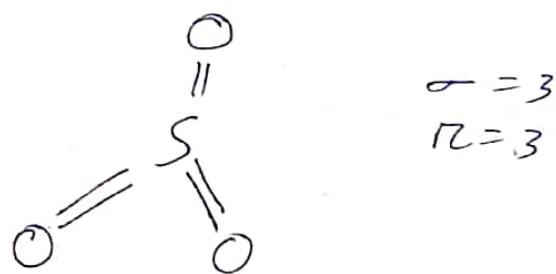
$$\sigma = 19$$

$$r = 9$$

$$\angle \rho = 4$$

$$\beta \cdot \rho = 18(9+9)$$

⑧ SO_3



$$\sigma = 3$$

$$r = 3$$

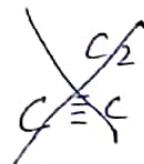
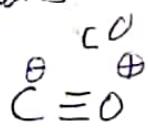
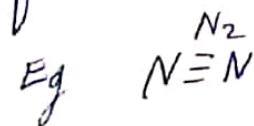
⑨ SO_2



$$\sigma = 2$$

$$r = 2$$

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

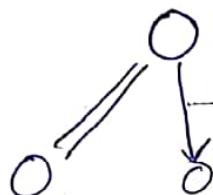


Types of Co-ordinate bonds

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

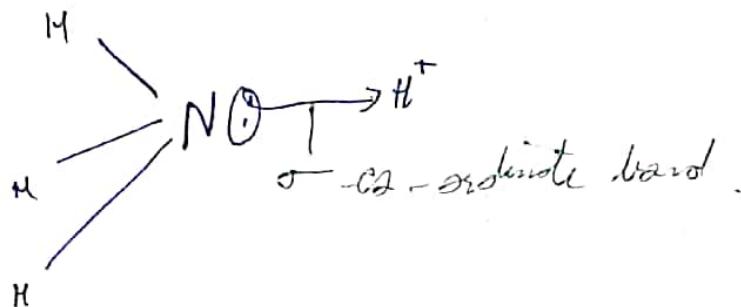
e.g.

O₃

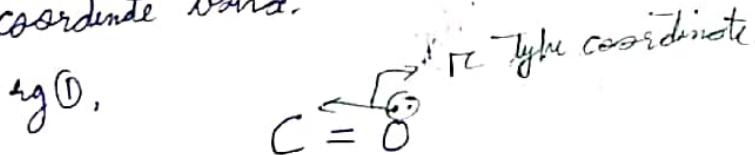


→ Type co-ordinate

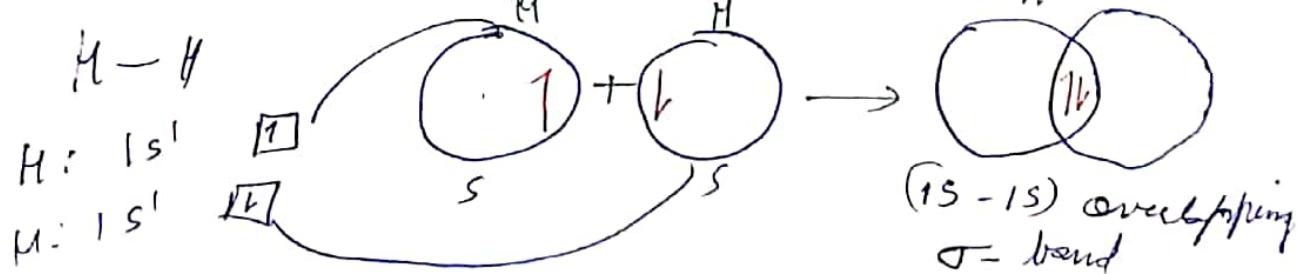
e.g(2)



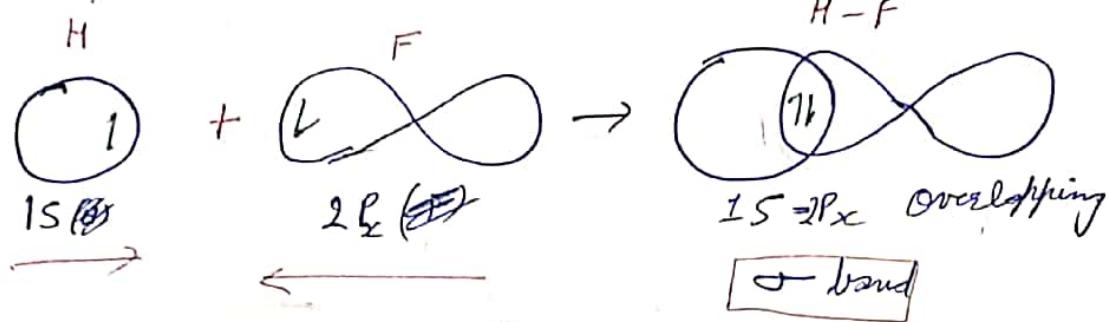
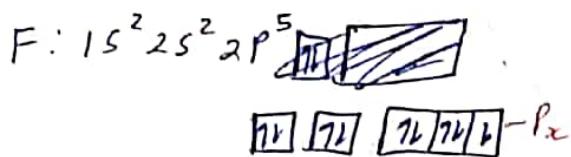
② π Type coordinate bond - If co-ordinate bond formed between 2 atoms is not first then it is known as π type coordinate bond.



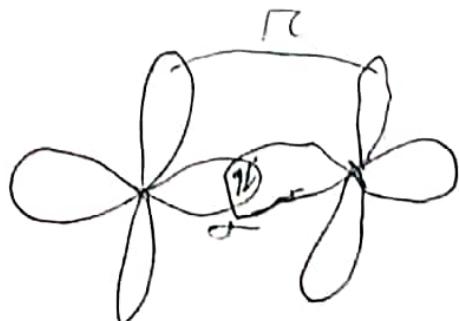
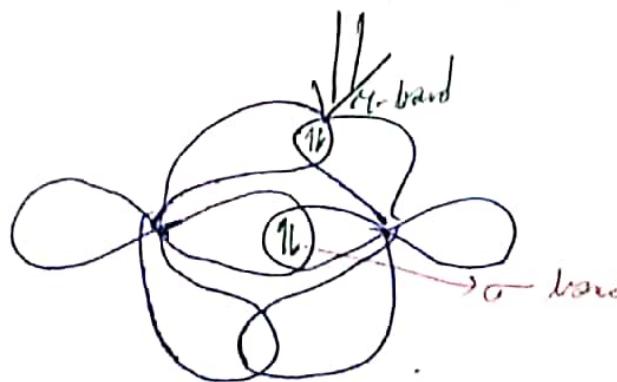
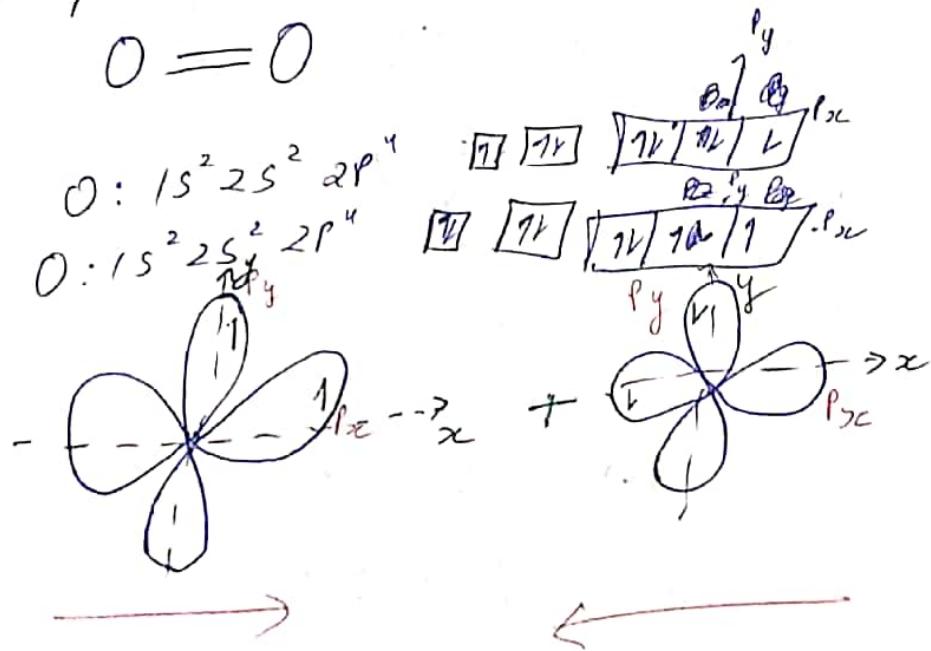
Q Explain H₂ molecule formation by VBT.



Q formation of HF explaining VBT



Q Explain O_2 formation using VBT.



Q explain N_2 formation by VBT?

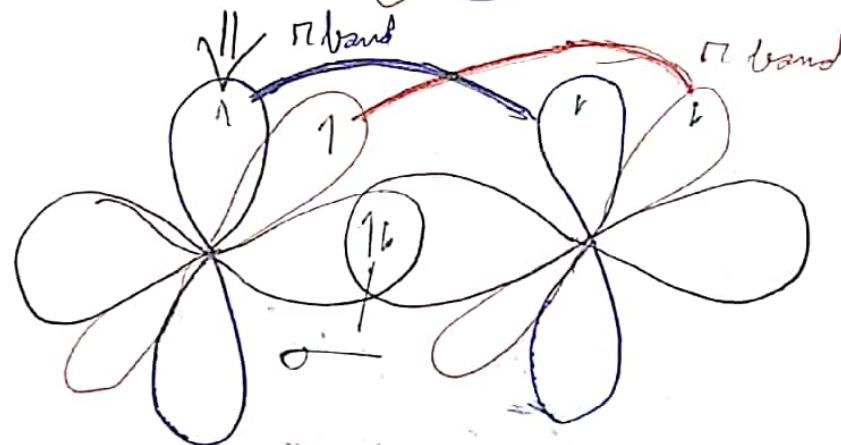
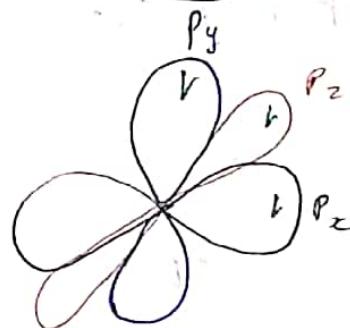
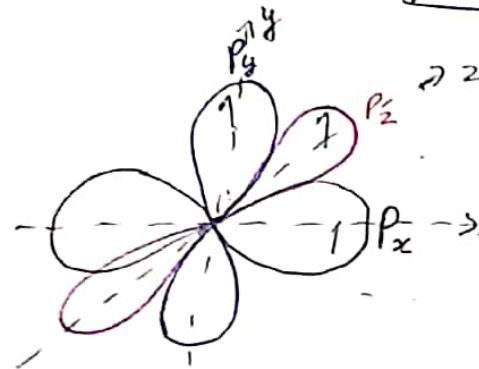
$N \equiv N$

$N: 1s^2 2s^2 2p^3$

$1s$	$1s$	p_x	p_y	p_z
1	1	1	1	1

$N: 1s^2 2s^2 2p^3$

$1s$	$1s$	p_x	p_y	p_z
1	1	1	1	1



(2p-2p) overlapping

σ -bond

- ① axial overlapping
- ② Stronger
- ③ Can Exist alone
- ④ Rotation around σ bond is possible

π -bond

- ①② Co-lateral overlapping
- ② Comparatively weaker
- ③ Cannot exist alone
- ④ ~~rotation around~~ Restricted Rotation

Strength of σ bonds :-

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

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

$(1s-1s) > (2s-2s) > (3s-3s)$ ~~as bond strength~~

$(2p-2p) > (3p-3p) > (4p-4p)$

$(1s-2p) > (1s-3p) > (1s-4p)$

Case 2 when n is same

Bond strength \propto directional nature of orbital.

directional nature $\Rightarrow s < p < d < f$
 n . dominates over directional nature

Q. Compare bond for σ -bond

① $(2s-2s) < (2s-2p) < (2p-2p)$

② $(1s-1s) \quad (1s-2s) \quad (2s-2p) \quad (2p-2p) \quad (2s-3s)$
 $(3s-3p) \quad (3p-3p) \quad (3s-4p) \quad (4s-4p) \quad (3s-4s)$

$(4p-4p) \quad (4s-4s)$

$(1s-1s) > (1s-2s) > (2p-2p) > (2s-2p) > (2s-3s) >$
 $(3p-3p) > (3s-3p) > (3s-4p) > (3s-4s) > (1p-4p) > (4s-4p) >$
 $(4s-4s)$

Strength of N₂ bond

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

Case 2 Bond strength of directional nature.

Eg ① $3P_{N_2} - 3P_{N_2}$ $2P_{N_2} - 2P_{N_2}$ $2P_{N_2} - 3P_{N_2}$

$$2P_{N_2} - 2P_{N_2} > 2P_{N_2} - 3P_{N_2} > 3P_{N_2} - 3P_{N_2}$$

② $2P_{N_2} - 2P_{N_2} > 3P_{N_2} - 3P_{N_2} > 4P_{N_2} - 4P_{N_2}$

③ $3P_{N_2} - 3P_{N_2} < 3P_{N_2} - 3d_{N_2} < 3d_{N_2} - 3d_{N_2}$

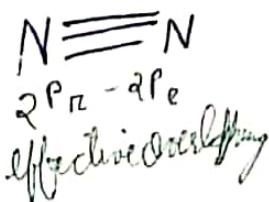
④ $2P_{N_2} - 3P_{N_2}$ $2P_{N_2} - 3d_{N_2}$ $2P_{N_2} - 4P_{N_2}$

$$2P_{N_2} - 3d_{N_2} > 2P_{N_2} - 3P_{N_2} > 2P_{N_2} - 4P_{N_2}$$

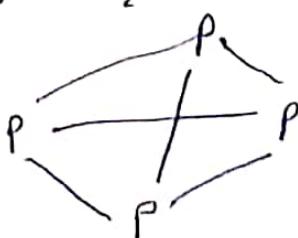
⑤ $2P_{N_2} - 2P_{N_2}$ $2P_{N_2} - 3d_{N_2}$ $2P_{N_2} - 3P_{N_2}$ $3P_{N_2} - 3P_{N_2}$

$$\underbrace{2P_{N_2} - 2P_{N_2}}_{\substack{\text{more stable} \\ (\text{effective overlapping})}} > \underbrace{2P_{N_2} - 3d_{N_2}}_{\substack{\text{less stable} \\ (\text{effective overlapping})}} > \underbrace{2P_{N_2} - 3P_{N_2}}_{\substack{\text{very less stable} \\ (\text{not effective overlapping})}} > \underbrace{3P_{N_2} - 3P_{N_2}}_{\substack{\text{Highly unstable} \\ (\text{not effective overlapping})}}$$

Q At room temp N₂ molecule exist but P₂ does not exist?



$P \equiv P$
 $3P_{N_2} - 3P_{N_2}$
not effective so not effective
instead exist as

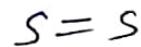


Q At Room Temperature O_2 exist but S_2 does not why.



$$2P_n - 2P_n$$

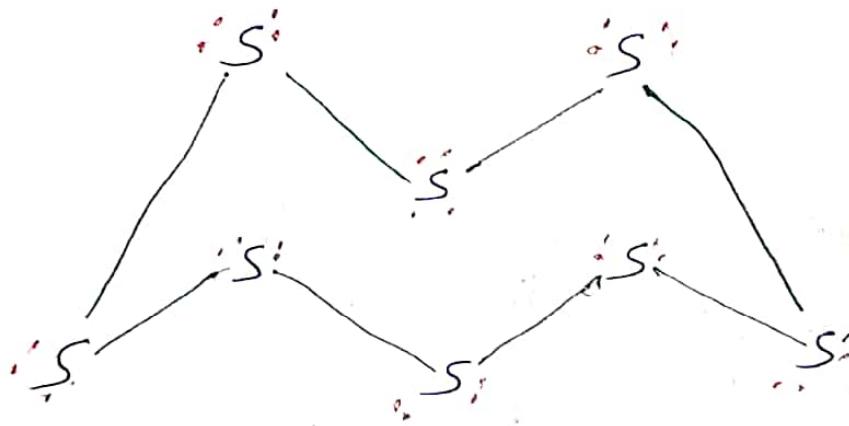
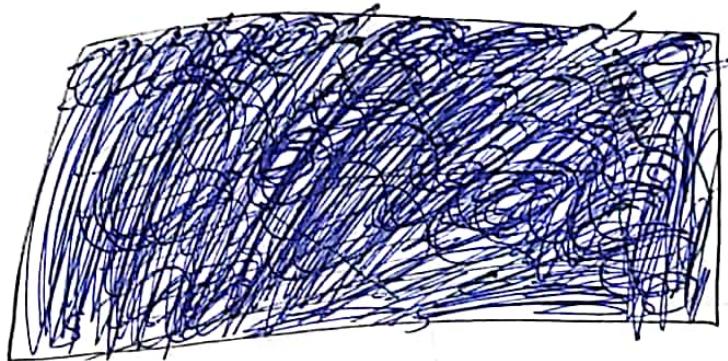
effective overlapping



$$3P_n - 3P_n$$

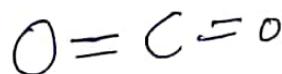
not effective

So exist as S_8 (crown like structure)



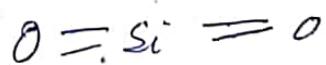
$$L \cdot P = 10$$

Q At room temperature C_2 exist in monomeric form but SiO_2 molecule exist in polymeric form.



$$(2P_n - 2P_n)$$

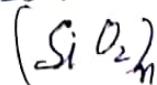
effective

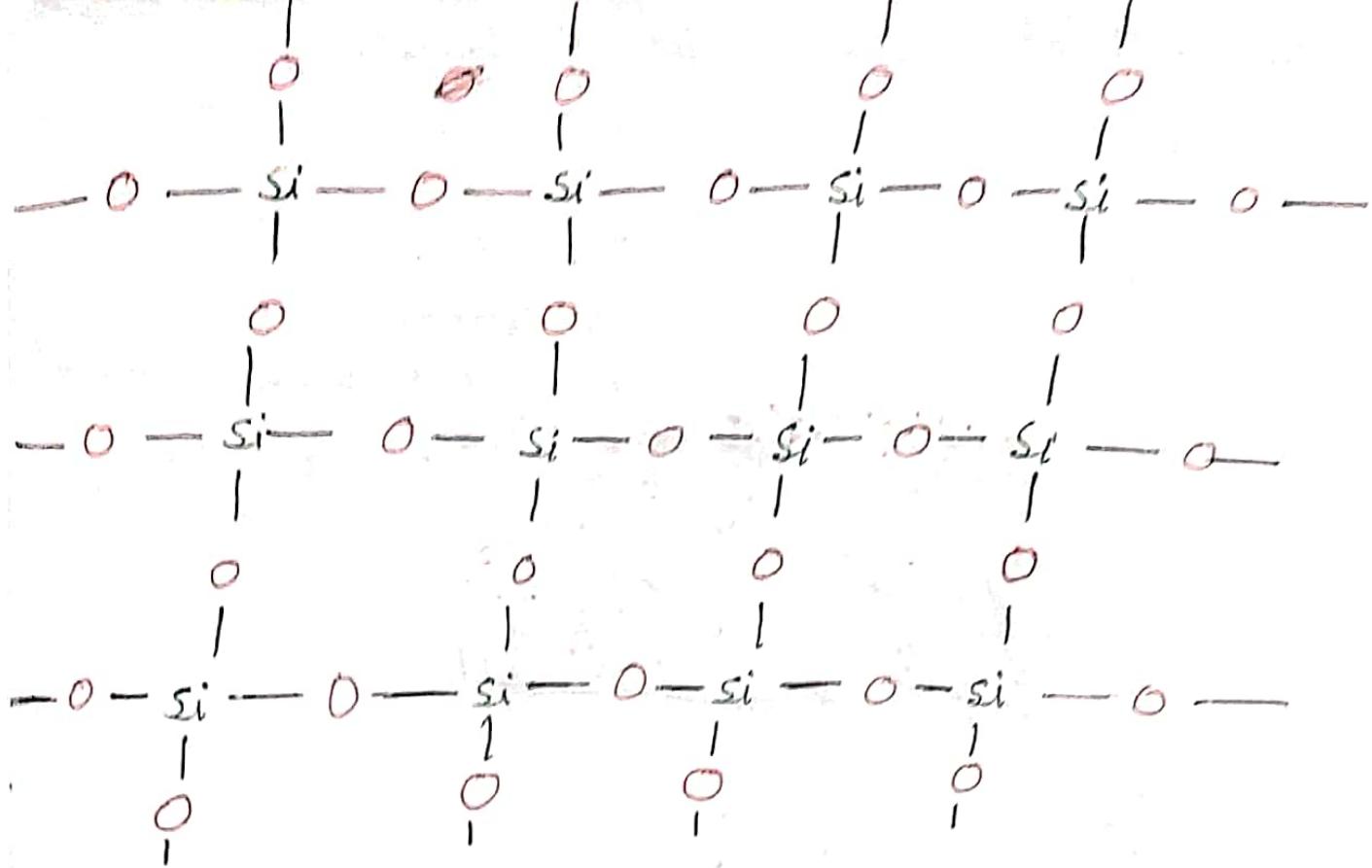


$$(2P_n - 3P_n)$$

unstable

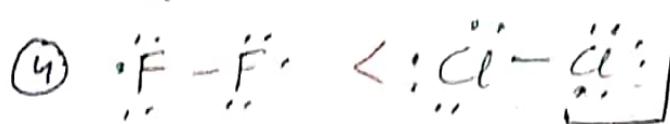
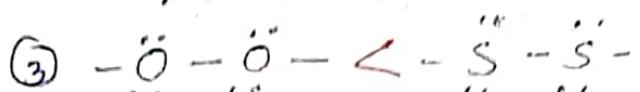
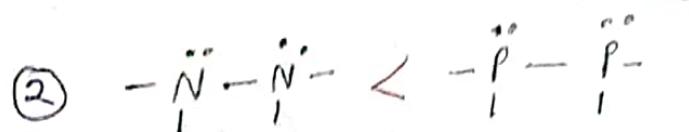
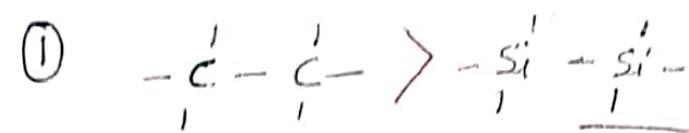
so exist in polymeric form





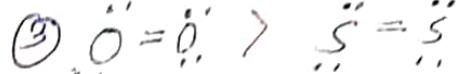
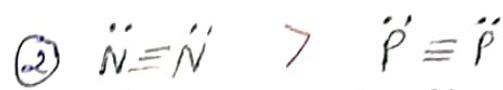
Bond Strength Orderz

Single Bond

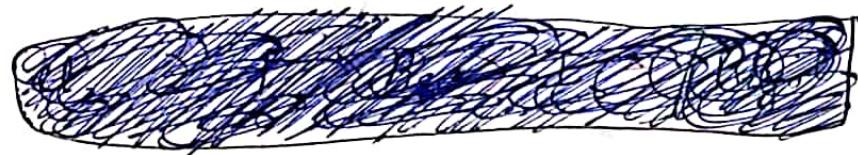


↓
due to small size L.P.
repulsion dominate over n

Multiple Bond



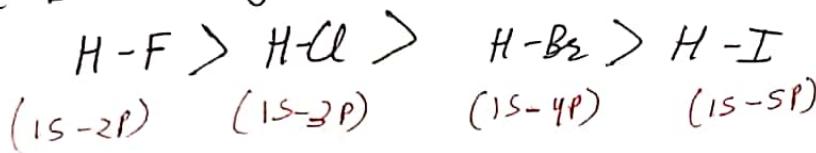
Q Bond strength order of F_2 , Cl_2 , Br_2 & I_2



expected: $F_2 > Cl_2 > Br_2 > I_2$

Real: $Cl_2 > Br_2 > F_2 > I_2$

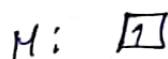
Q Bond strength



Drawbacks of VBT

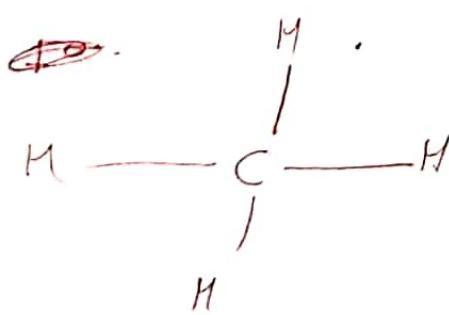
① Valence Bond Theory fails to explain geometry and bonding of molecules containing more than 2 atoms.

① CH_4



1σ band: 1s-2s (weaker)

3σ band: 1s-2p (stronger)



→ According to VBT CH_4 molecule has no all

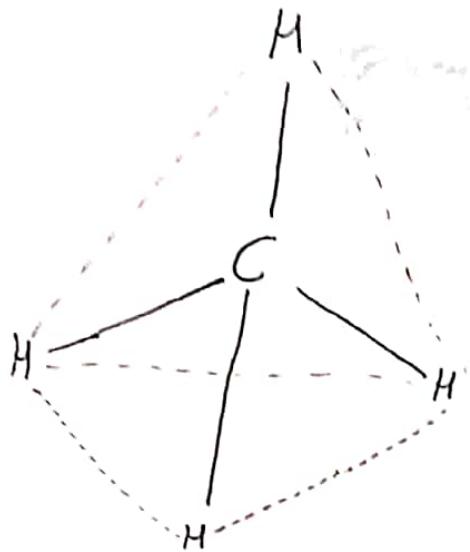
4 (C-H) bond ~~length~~ length & bond strength equal

→ But in real all 4 (C-H) bond lengths and bond strengths are some & 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 ~~weak~~
but in reality all the 4 bonds have ~~all~~ some bond length.

according to VBT

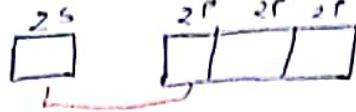


Hybridisation

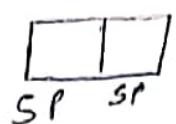
→ It is the best theoretical concept to explain geometry of molecules.

Assumptions -

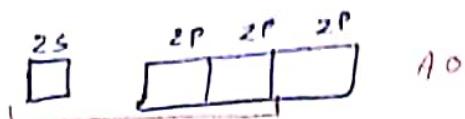
- ① Hybridisation is the ~~for~~ intermixing of atomic orbitals which has comparable energy or less energy difference.
- ② As many no. of atomic orbitals are participated, some no. of hybrid orbitals will be formed.
- ③ Once the hybrid orbitals are formed they arrange themselves in such a way so that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



Atomic Orbitals (A.O.)

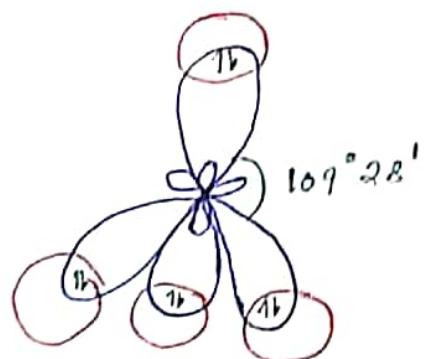
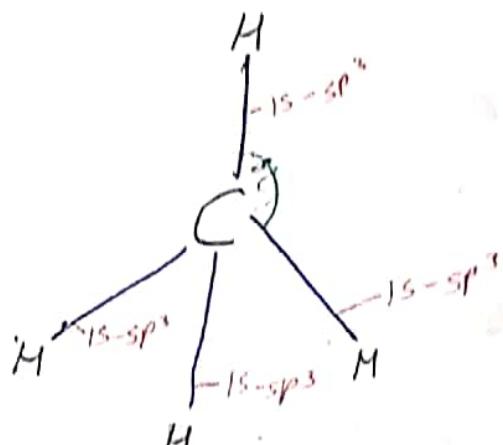
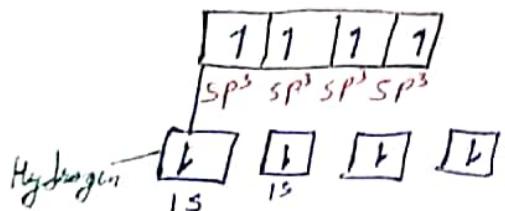
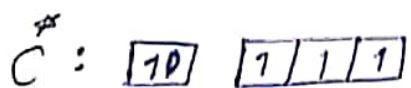


Hybrid orbital (H-O)



H-O

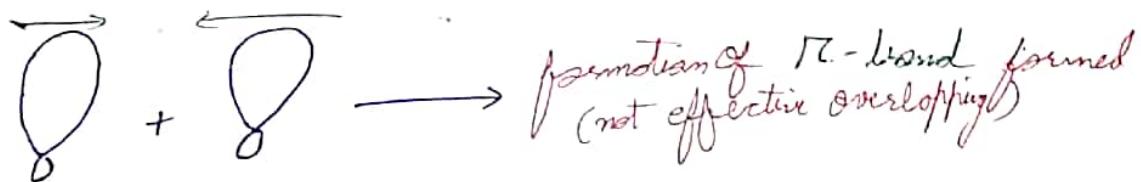
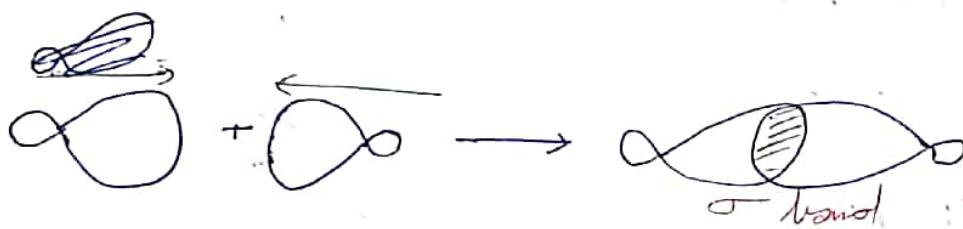
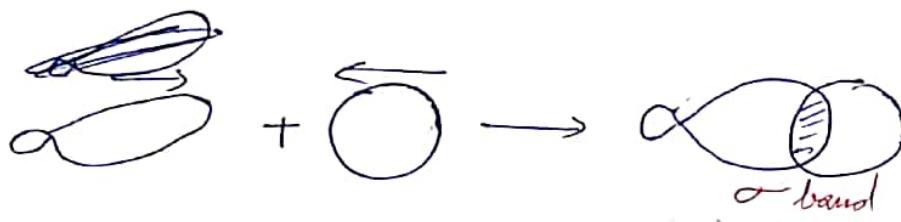
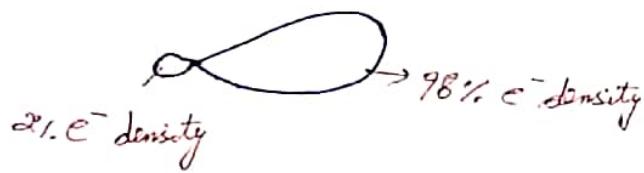
Ex ①



- ② → Here 4 (C-H) bonds are formed \Rightarrow This explain equal bond length, bond energy and bond strength.
- In CH₄, all bonds formed are identical with overlapping of (SP³-1s)

(3D)

Shape of Hybrid Orbital



* 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 V.P character $\propto \frac{1}{\% S \text{ character}}$

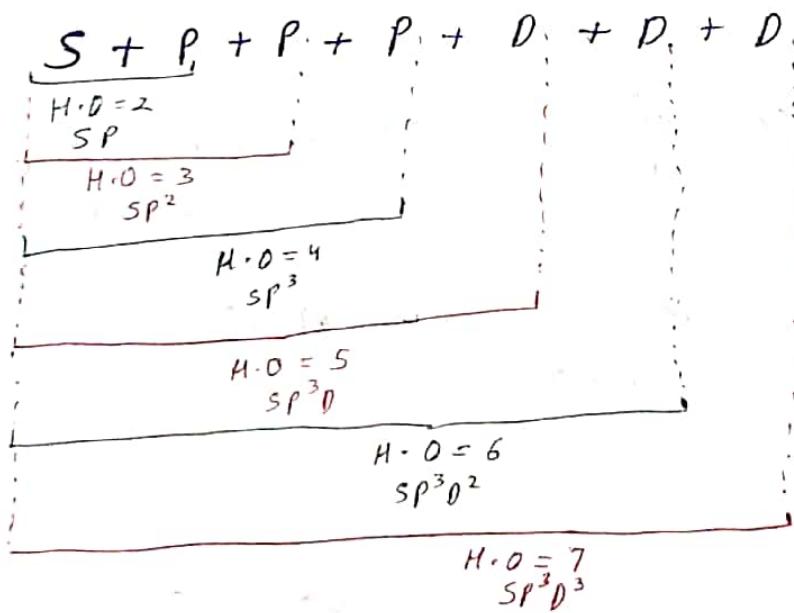
$SP < SP^2 < SP^3 \}$ Energy of H-O \propto $\frac{\% P \text{ character}}{\% S \text{ character}}$

$SP > SP^2 > SP^3 \}$ stability of H-O \propto V.S character

Electron Negativity \propto V.S character

$$\text{EN: } SP > SP^2 > SP^3$$

Calculation of Hybridisation



Steric no. (S.N.) \Rightarrow Total no. of hybrid orbital

\Rightarrow no. of σ - bond + no. of L.P. on central atom
of central atom

\Rightarrow no. of surrounding atom + no. of L.P. on central atom.

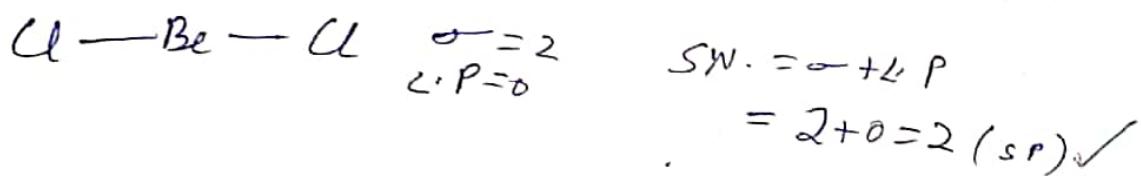
$SN -$

Type of Hybridisation

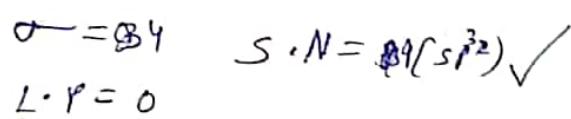
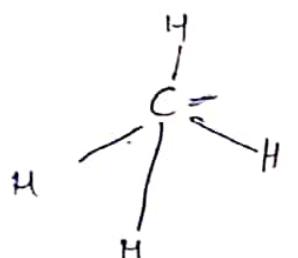
2	SP
3	SP^2
4	SP^3
5	SP^3D
6	SP^3D^2
7	SP^3D^3

Q Find the hybridisation of central atom?

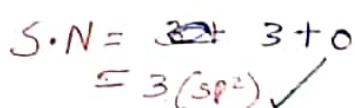
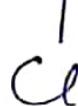
① $BeCl_2$



② CH_4

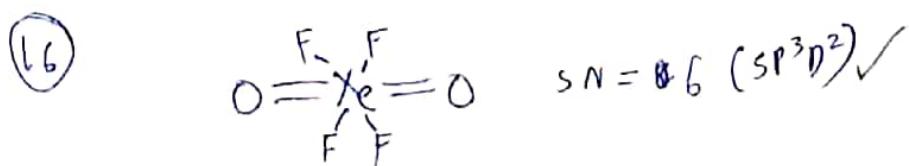
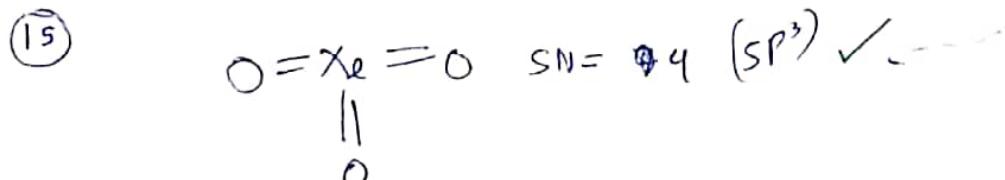
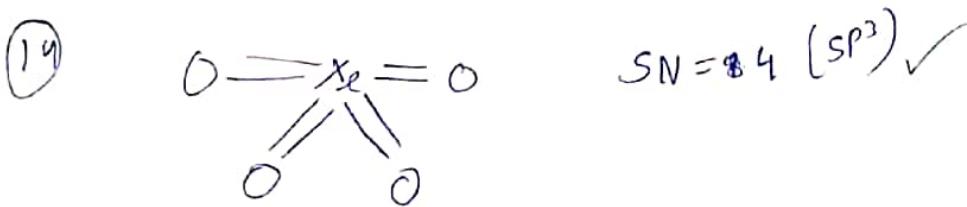
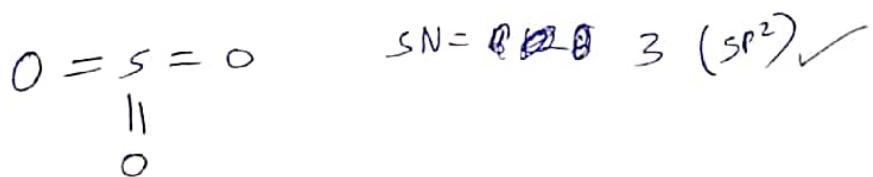
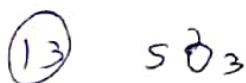
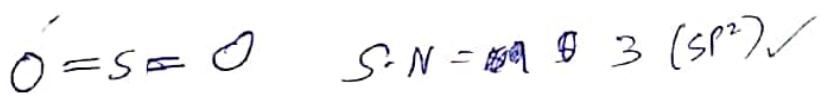
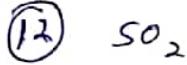
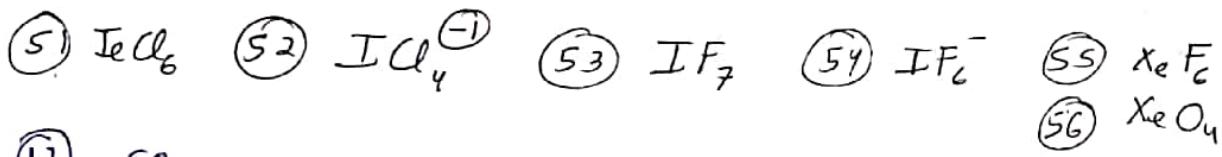


③ $AlCl_3$ $Cl - Al - Cl$

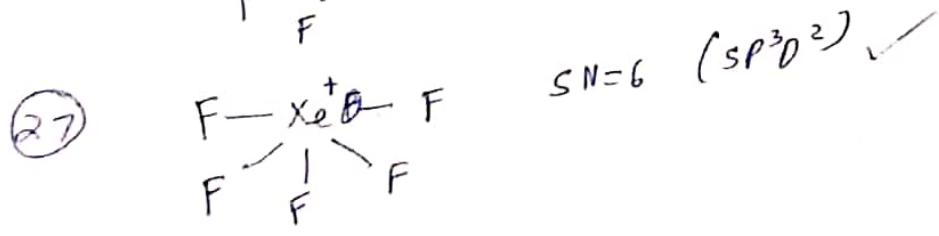
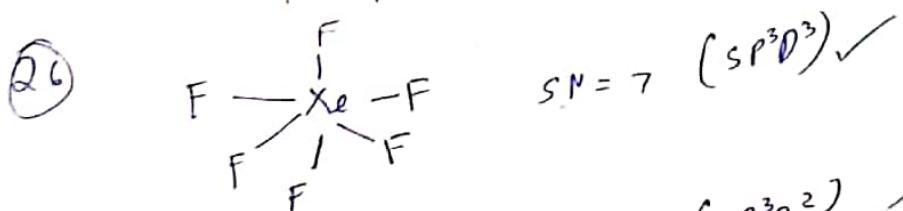
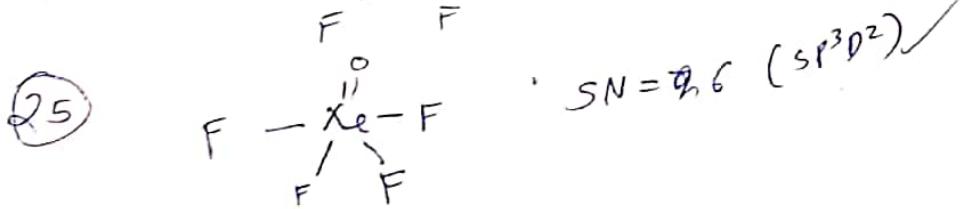
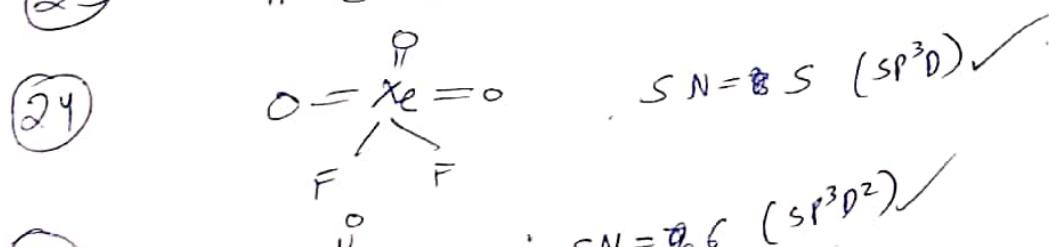
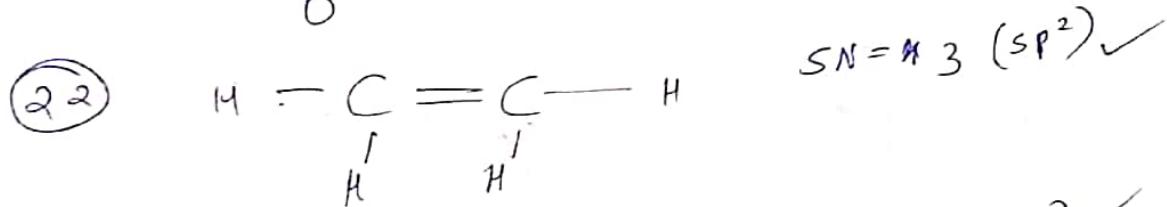
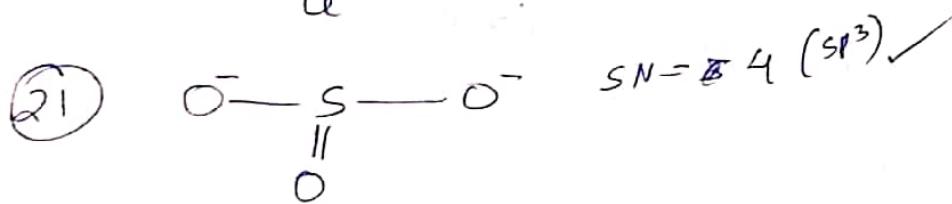
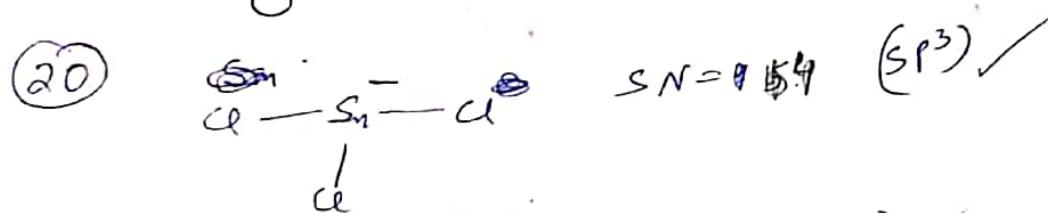
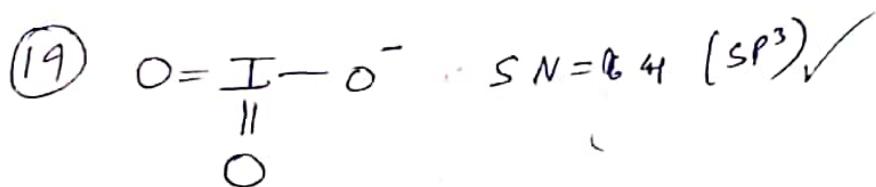
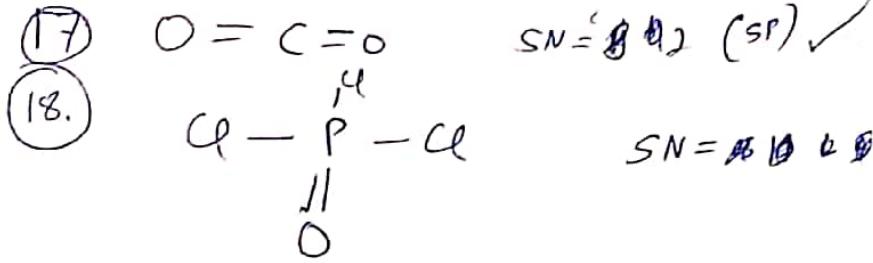


- ④ $\text{P} \text{Cl}_5$ ④ ~~PBr_3~~ ⑤ $\text{N} \text{H}_3$ ⑥ H_2O ⑦ SF_6 ⑧ IF_7 ⑨ SCl_2 ⑩ XeF_2 ⑪ XeF_4
- $SN = 5 (sp^3d)$ ✓
-
- ⑤ NH_3 ⑥ H_2O ⑦ SCl_2 ⑧ IF_7 ⑨ SCl_2 ⑩ XeF_2 ⑪ XeF_4
- $SN = 3 + 1 = 4 (sp^3)$ ✓
-
- ⑥ $\text{H} - \text{O} - \text{H}$ ⑦ $\text{F} - \text{S} - \text{F}$ ⑧ $\text{F} - \text{I} - \text{F}$ ⑨ $\text{Cl} - \text{S} - \text{Cl}$ ⑩ ~~NO_2~~ - ~~SF_4~~
- $SN = 2 + 2 = 4 (sp^3)$ ✓
- $SN = 6 (sp^3d^2)$ ✓
- $SN = 7 (sp^3d^3)$ ✓
- ⑪ ~~XeF_4~~
- $SN = 2 + 2 = 4 (sp^3)$ ✓
- $SN = 2 + 3 = 5 (sp^3d)$ ✓
- $SN = 4 + 2 = 6 (sp^3d^2)$ ✓
- (43)

- (12) SO_2 (21) SO_3^{2-} (30) XeF_5^+ (40) ClO_4^{-1}
 (13) SO_3 (22) C_2H_4 (31) IF_{47}^+ (41) H_3O^+
 (14) XeO_4 (23) C_2H_2 (32) IF_6^- (42) SF_4
 (15) XeO_3 (24) XeO_3F_2 (33) I_3^+ (43) SeF_4
 (16) ~~(1)~~ XeO_2F_4 (25) $XeOF_4$ (34) C_2H_6 (44) $TeCl_4$
 (17) CO_2 (26) XeF_6 (35) BF_2 (45) ClF_3
 (18) $POCl_3$ (27) XeF_5^+ (36) ~~(1)~~ SO_3 (46) ICl_2^-1
 (19) IO_3^- (28) NO_2^+ (37) O_3 (47) Br_2F_3
 (20) $SnCl_3^-$ (29) NO_2^- (38) NH_4^+ (48) $Br_2F_2^-$
 (39) PH_4^+ (49) XeO_6^{4-}
 (50) $Br_2F_4^{-1}$

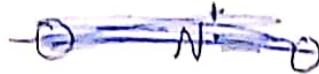


(44)

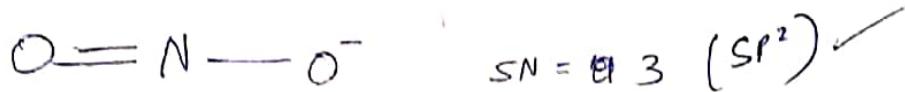


(43)

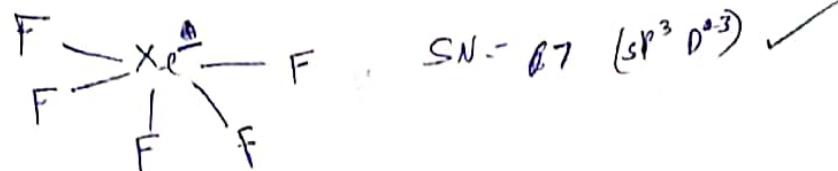
(28)



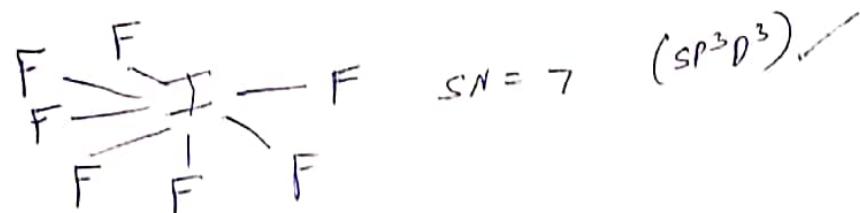
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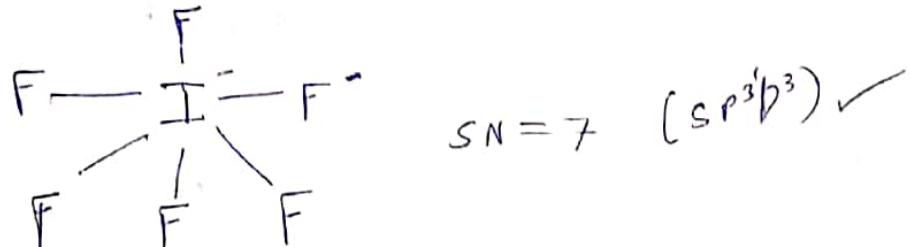
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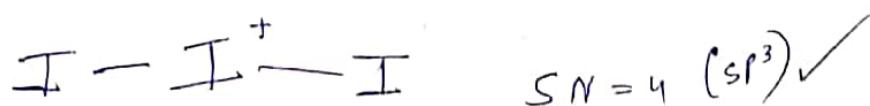
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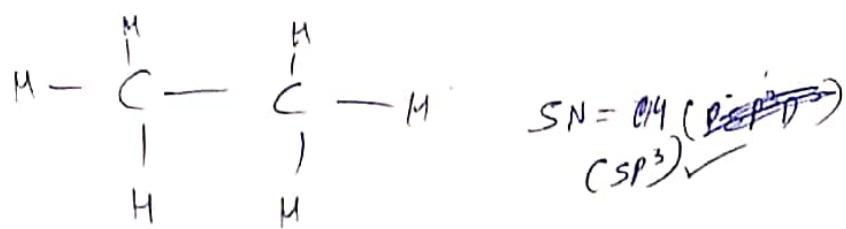
(32)



(33)



(34)

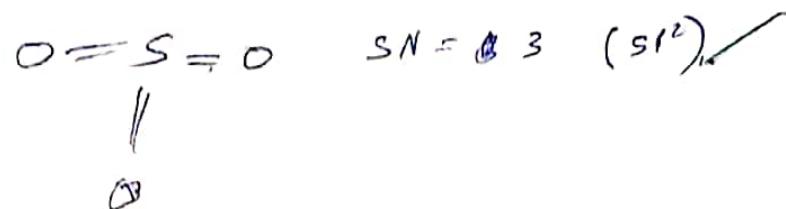


(34.5)

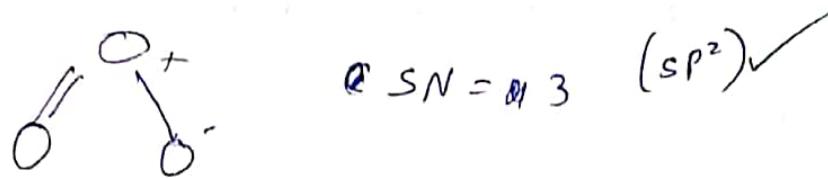


(46)

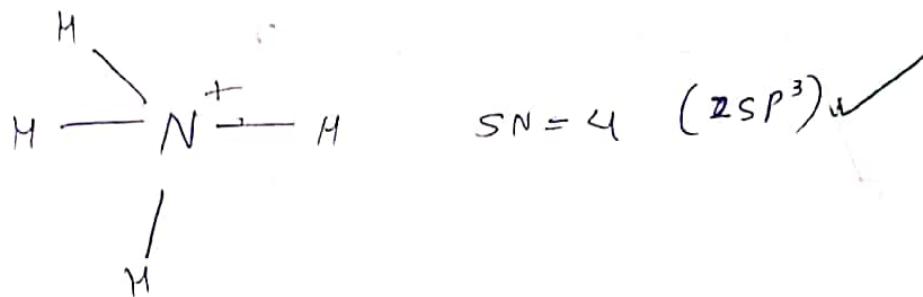
(36)



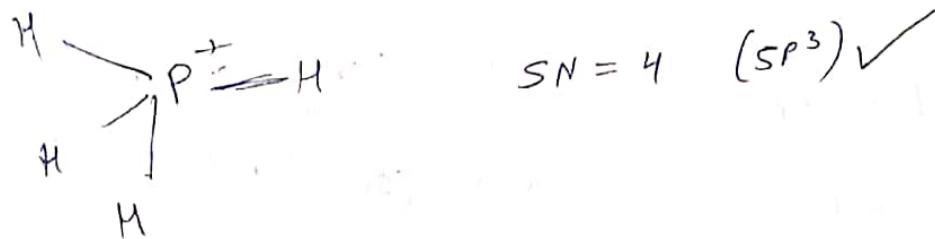
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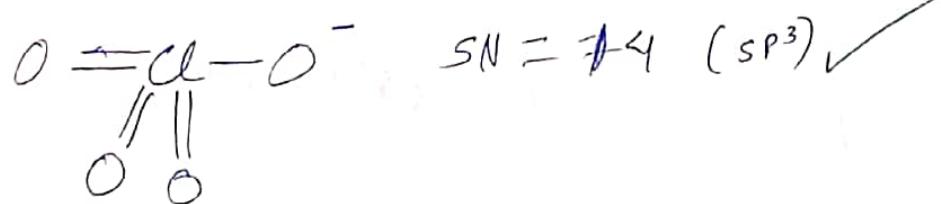
(38)



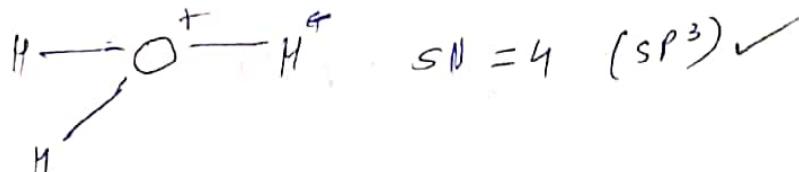
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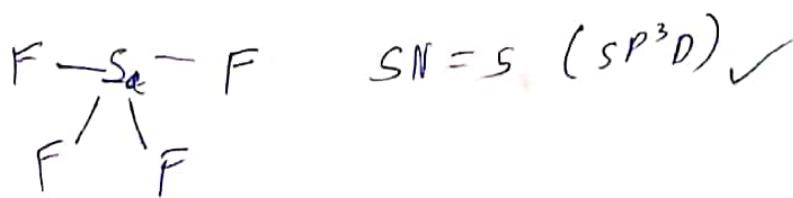
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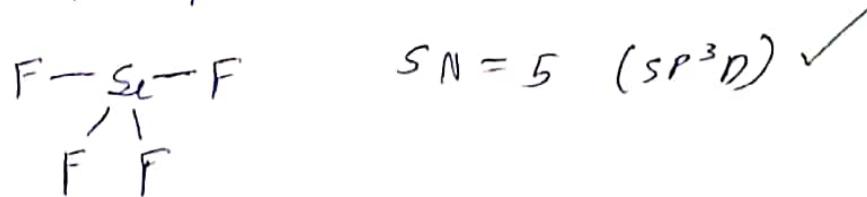
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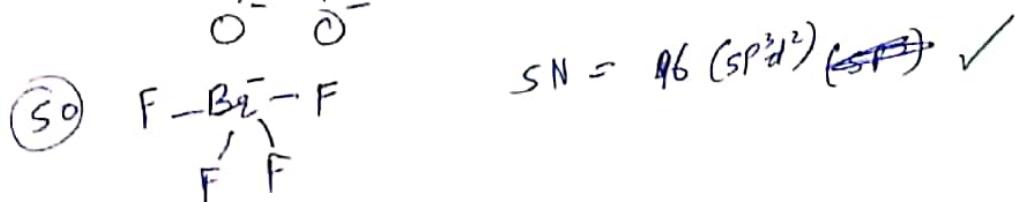
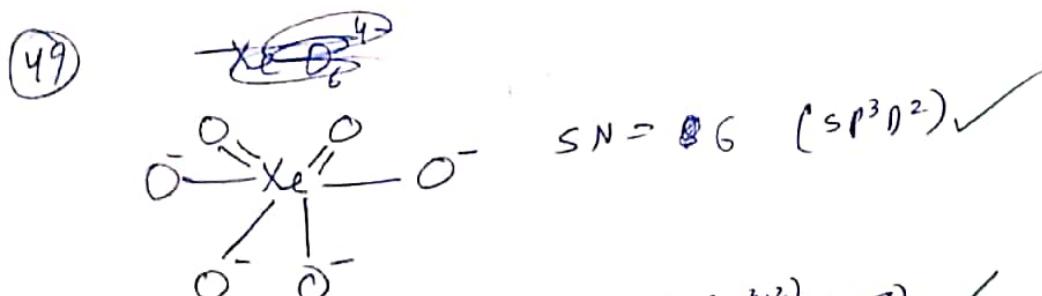
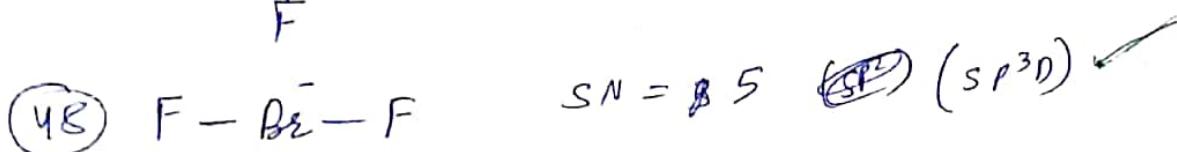
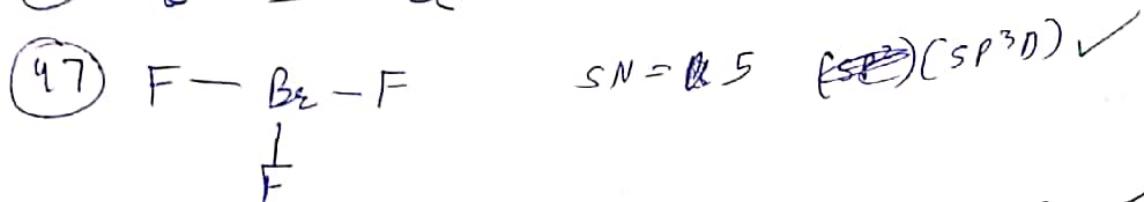
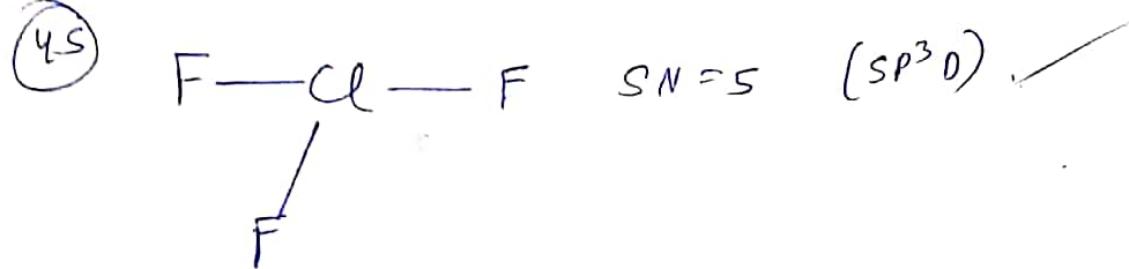
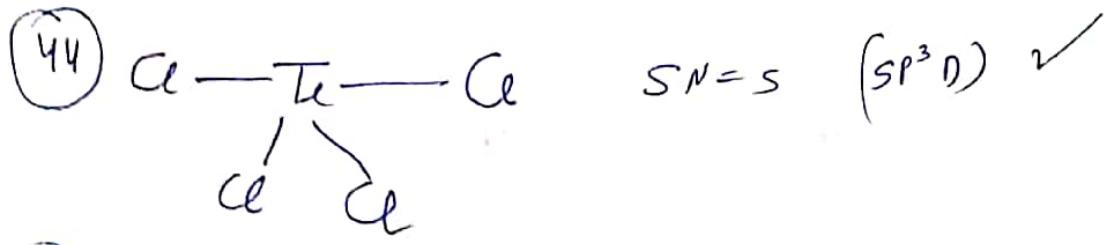
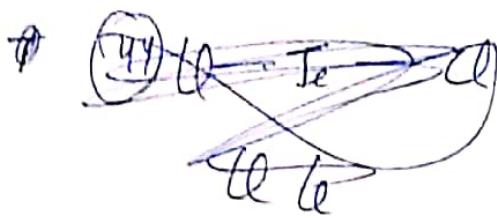
(42)



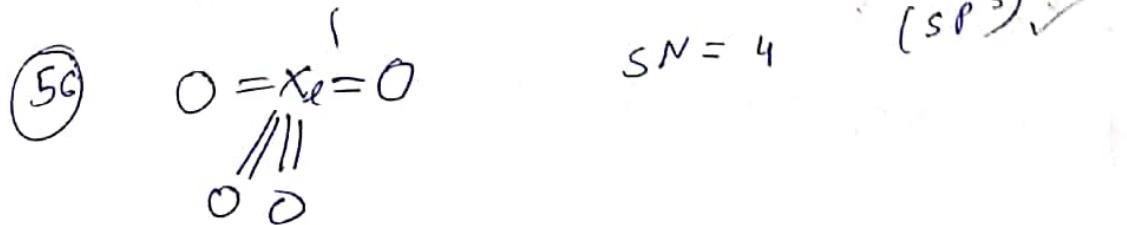
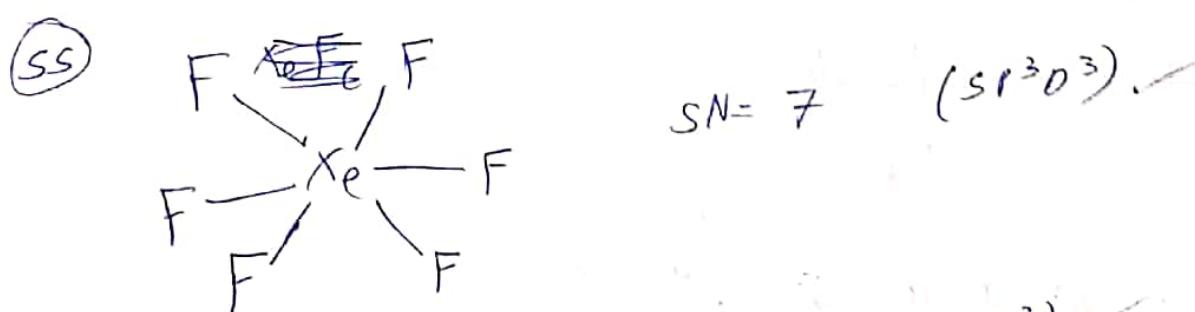
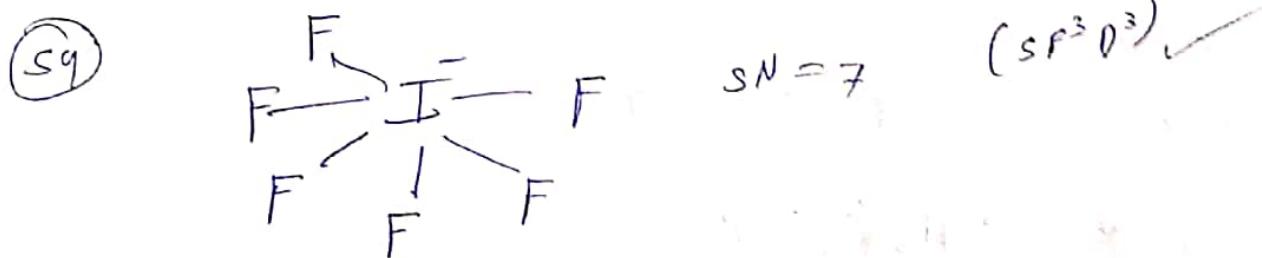
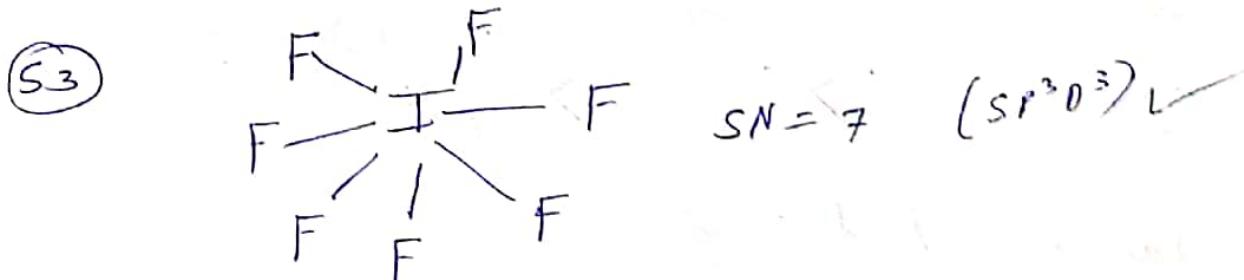
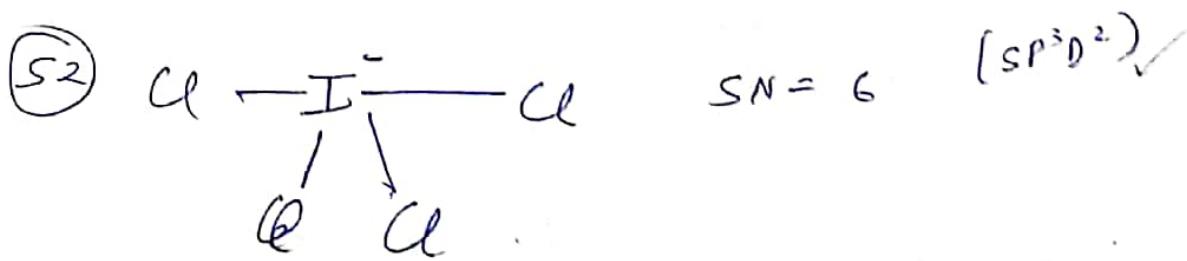
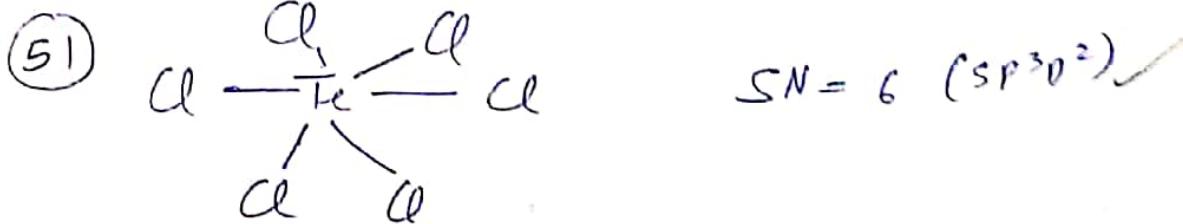
(43)



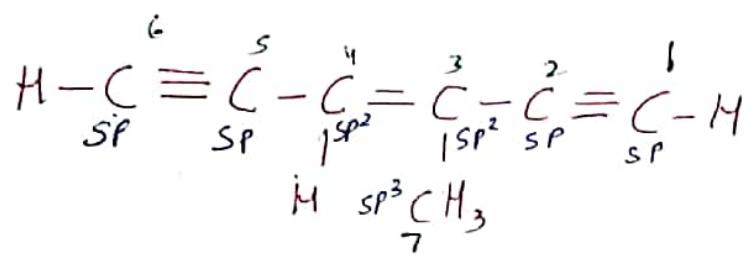
(47)



48



Q EN order?

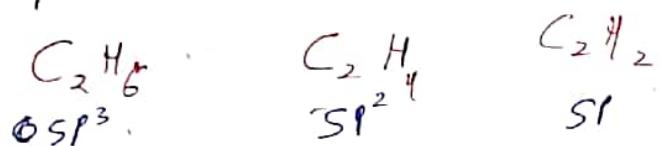


EN & I.S. character

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

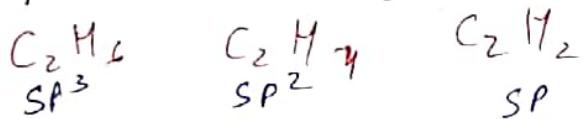
~~Electrostatic Resonance~~ $(C_1 = C_2 = C_s = C_d) > (C_4 = C_3) > C_7$ [EN order]

Q 2. Compare C-C bond strength



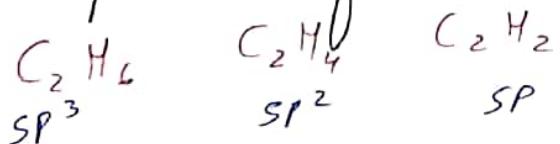
$C_2H_6 < C_2H_4 < C_2H_2$

Q 3. Compare C-C bond length



$C_2H_6 > C_2H_4 > C_2H_2$

Q 4. Compare E.N. of the central atom



$C_2H_6 < C_2H_4 < C_2H_2$

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

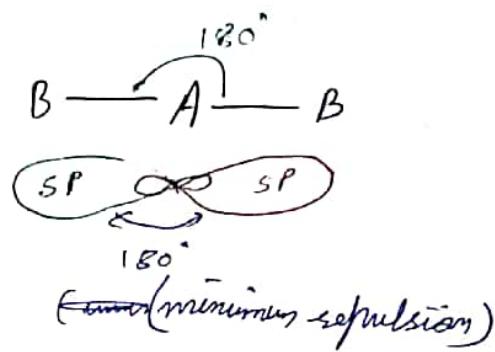
① SP-Hybridisation -

$S + (\text{only } p \text{ orbital})$

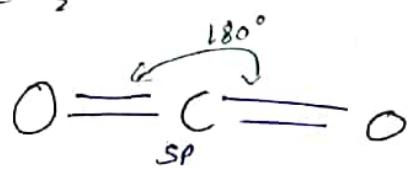
$S + (p_x/p_y/p_z)$

Geometry: linear (planar)

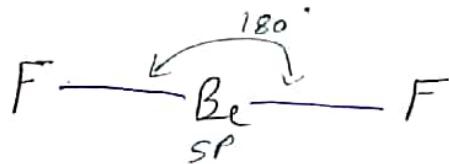
Angle: 180°



E.g. CO_2



E.g. ② BeF_2



② SP^2 -Hybridisation -

$S + (\text{only } 2 p \text{ orbitals})$

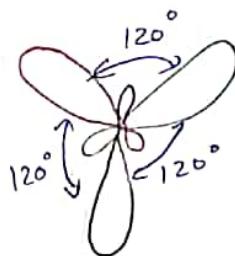
~~$S + (2 p_x/2 p_y/2 p_z)$~~

$S + p + p$

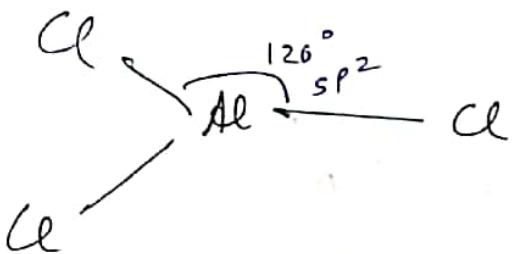
Geometry: Trigonal Planar

Bond angle: 120° (3)

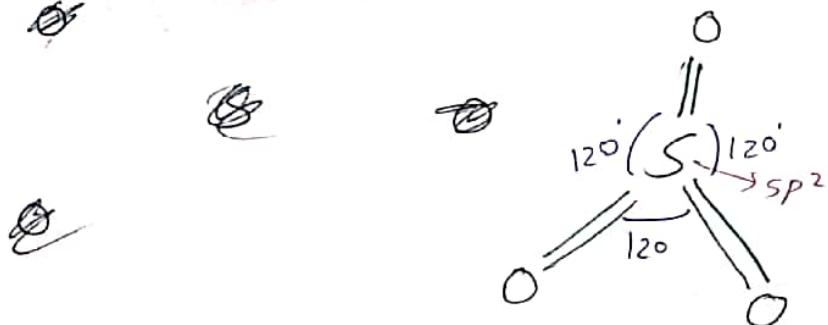
(Planar)



Eg ① AlCl_3



② SO_3



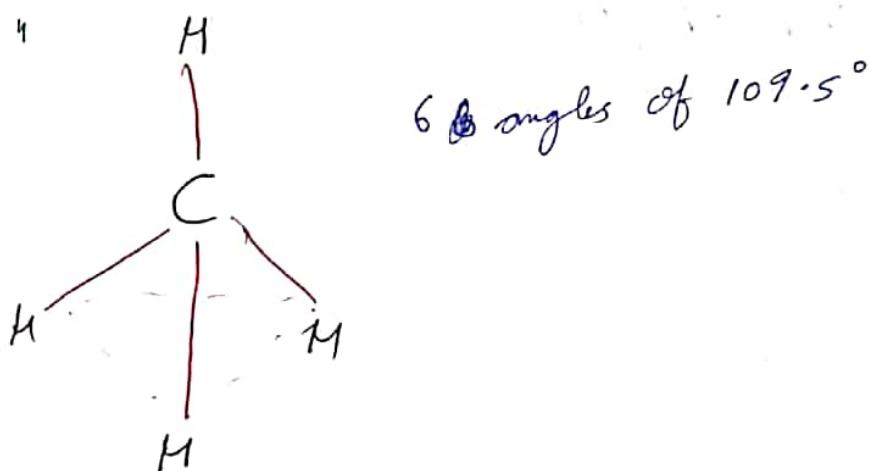
③ sp^3 -Hybridisation -

$$(s + p_x + p_y + p_z)$$

Geometry : Tetrahedral

Bond Angle : $109.5^\circ / 109^\circ 28'$
(non-Ca-Planer)

Eg ①. CH_4



④ $sp^3 d$ Hybridisation

$$S + P_x + P_y + P_z + d_{z^2}$$

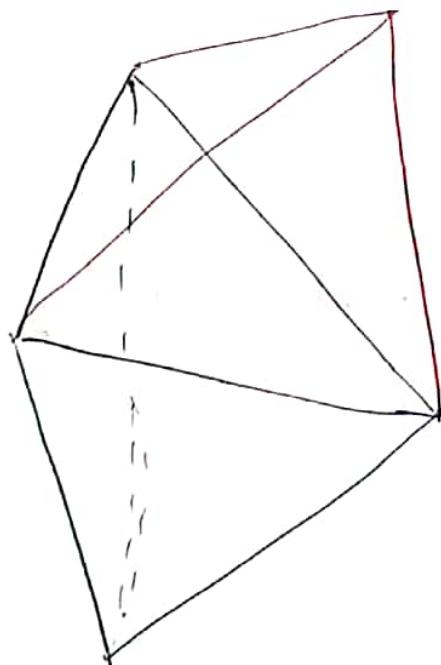
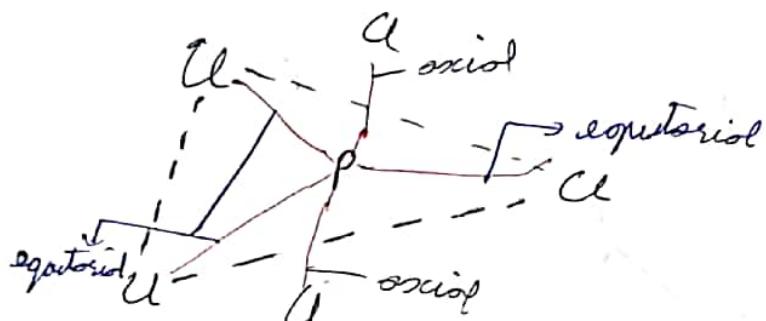
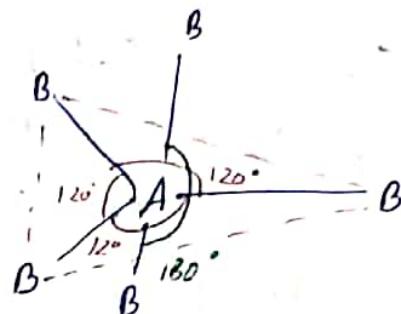
Geometry - Trigonal Bipyramidal

Bond Angl - 120° (3)

90° (6)

Non-Planar 180° (1)

Eg ① PCl_5



(5) sp^3d^2 Hybridisation

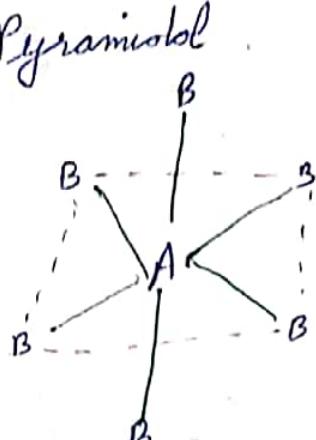
$$(s + p_x + p_y + p_z + d_{z^2} + d_{x^2-y^2})$$

Geometry :- Octahedral / Square Bi Pyramidal

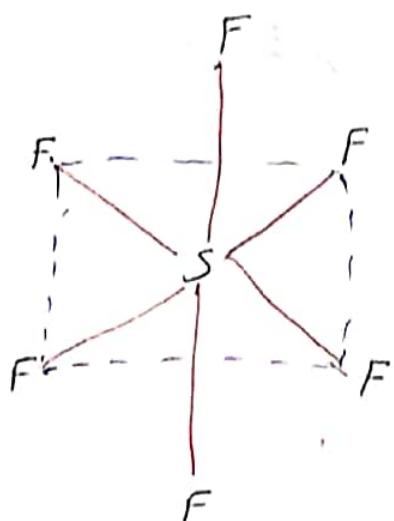
Bond Angle :- $90^\circ(12)$

$180^\circ(3)$

Non-Planar



Eg. SF_6



(6) sp^3d^3 Hybridisation

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

Geometry :- Pentagonal Bi Pyramidal

Bond Angle :- $180^\circ(1)$

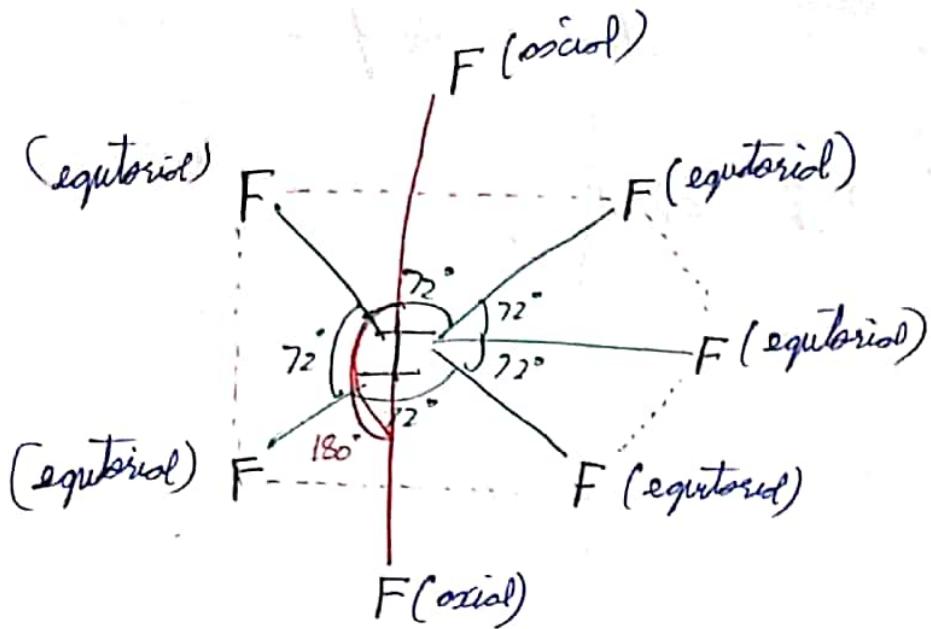
$90^\circ(10)$

$72^\circ(5)$

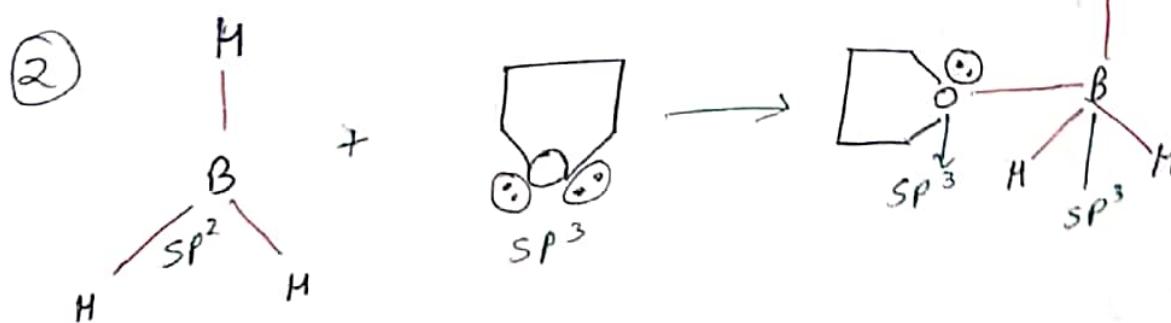
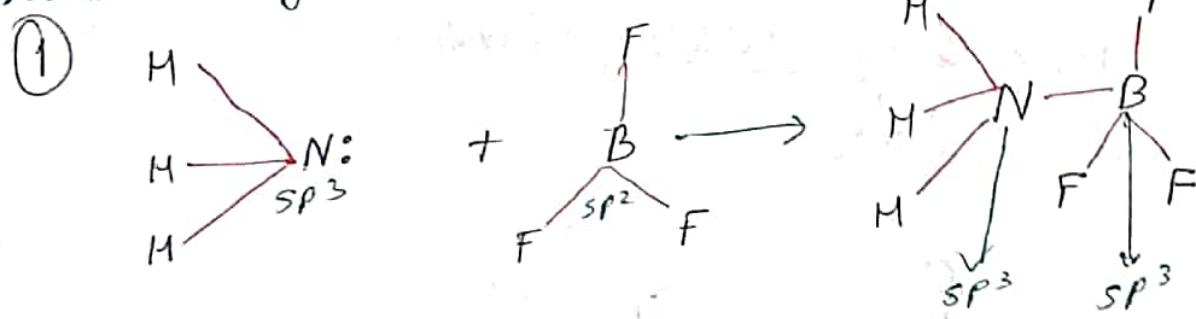
$144^\circ(5)$

Non-Planar

(54)



~~Expt~~ End Hybridisation



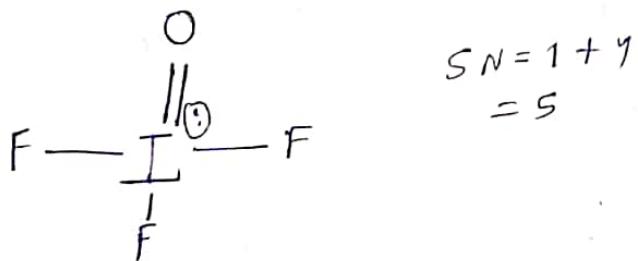
Q In ϕSP^3d hybridisation which of the following orbitals may be involved.

- 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 $nS + \text{only } 2\text{-nd} + 3\text{-rd } P \rightarrow SP^3d^2$
④ $nS + 3\text{-rd } P + 2(n-1)d \rightarrow SP^3d^2$
 ⑤ $nS + 3\text{-rd } P + 2 \text{ octet } (n-1)d \rightarrow SP^3d^2$

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



- a) s
 b) p_z
c) $d_{x^2-y^2}$
 d) d_{z^2}

Q Match the column:

specis

- A) SO_4^{2-}
B) SO_3 ($x=2$ -none)
C) XeO_2F_2
D) XeF_4
E) SF_2Cl_2

Toric orbital involved in hybridisation.

- P) s
Q) p_z
R) d_{xy}
S) $d_{x^2-y^2}$
T) d_{z^2}

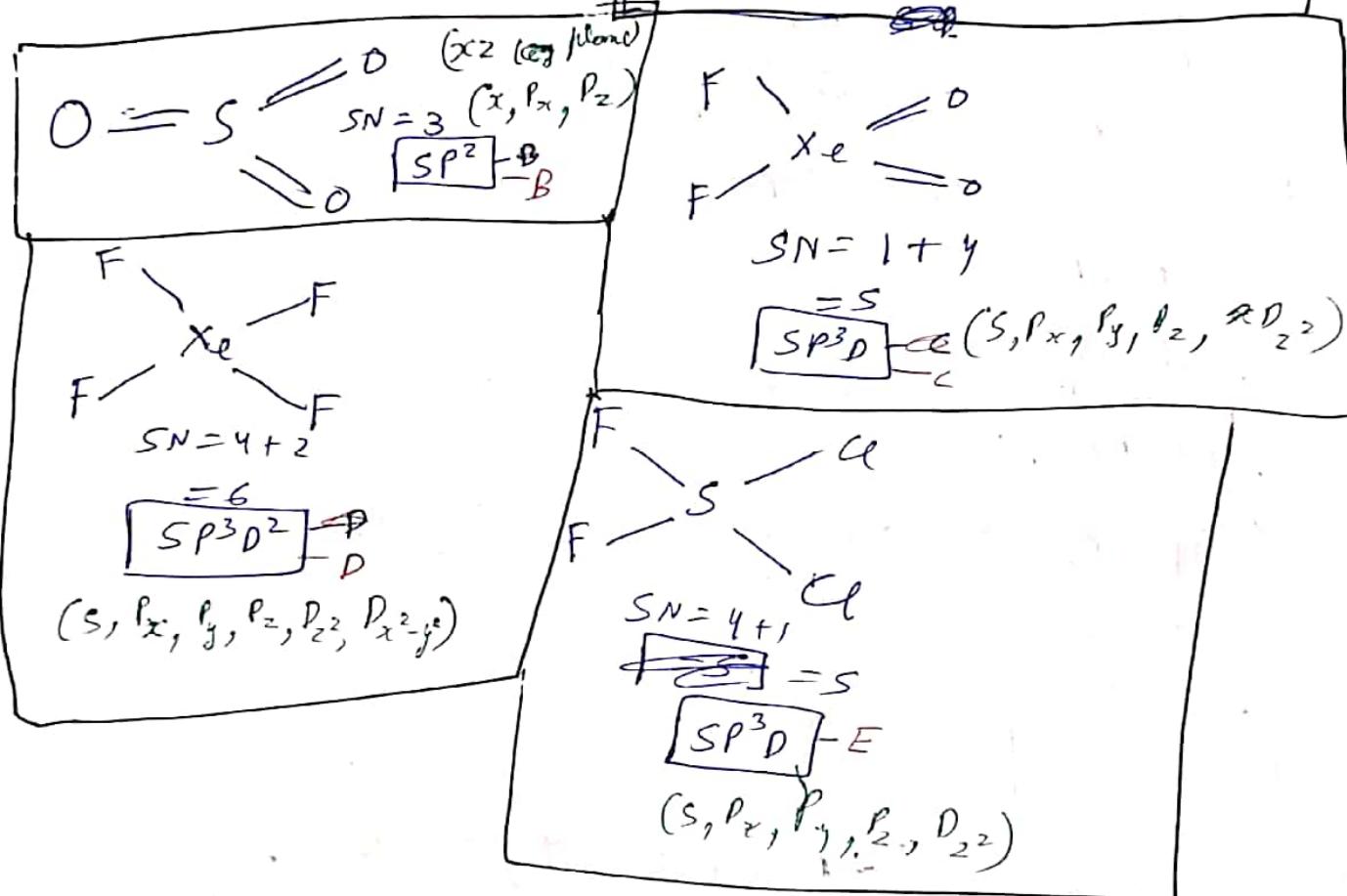
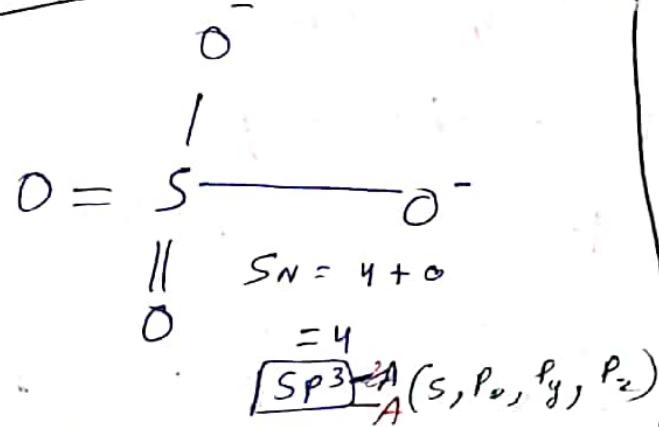
A - PQ

B - PQ

C - PQT

D - POTS

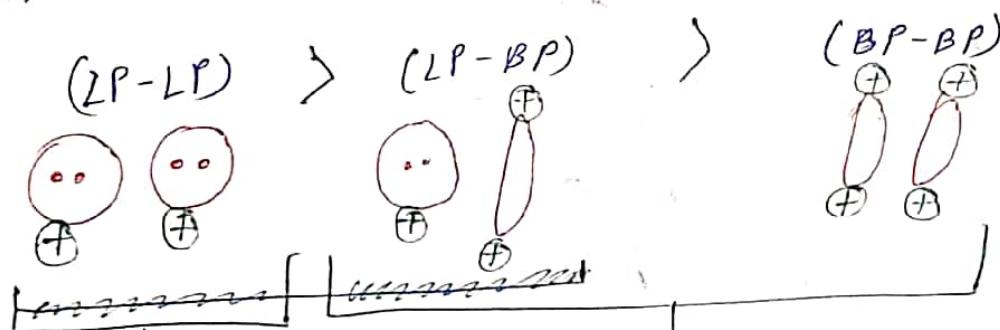
E - PQT



VSEPR THEORY

(Valence Shell e⁻ pair repulsion theory)

- According to VSEPR theory, Bond Pair & Lone Pairs repel each other. To minimize these repulsions, molecule occurs 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

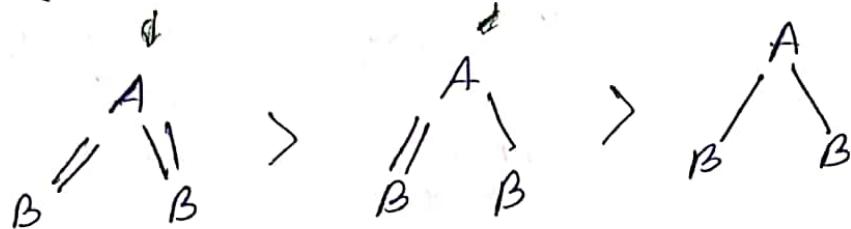


- Since the size of Lone Pair e⁻ cloud is larger due to absence of 2nd nucleus.
- Hence the distance b/w two lone pair is less & repulsion is high.
- Bond Pairs get flattened due to stretching by two nucleus. Hence the distance in (LP-BP) or (BP-BP) is increased & repulsion is lower.

Typhus of Bond Force Repulsion

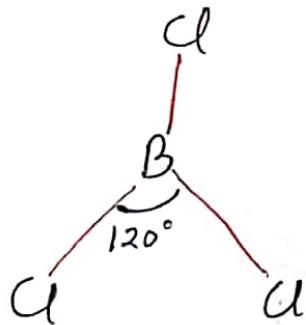
- ① Multiple Bond - Multiple Bond
- ② Multiple Bond - Single Bond
- ③ Single Bond - Single Bond

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



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

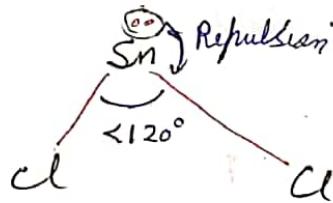
Eg ① SP^2



$$\sigma = 3$$

$$l \cdot P = 0$$

$$SN = 3$$

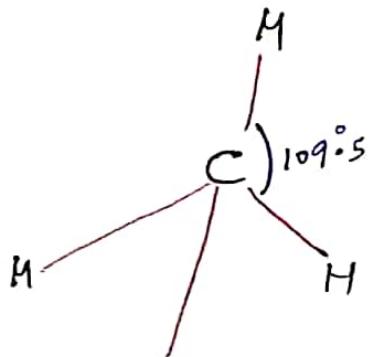


$$\sigma = 2$$

$$l \cdot P = 1$$

$$SN = 3$$

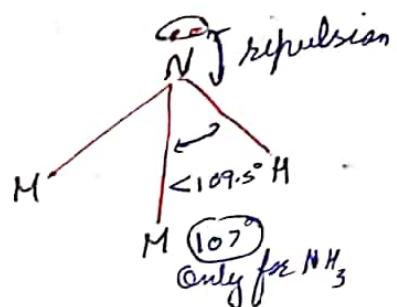
② SP^3



$$\sigma = 4$$

$$L.P = 0$$

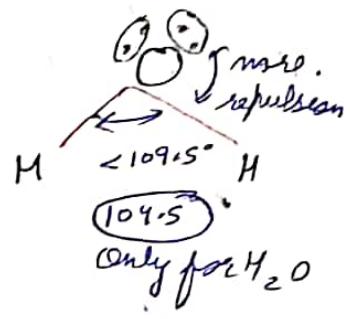
$$S.N = 4$$



$$\sigma = 3$$

$$L.P = 1$$

$$S.N = 4$$



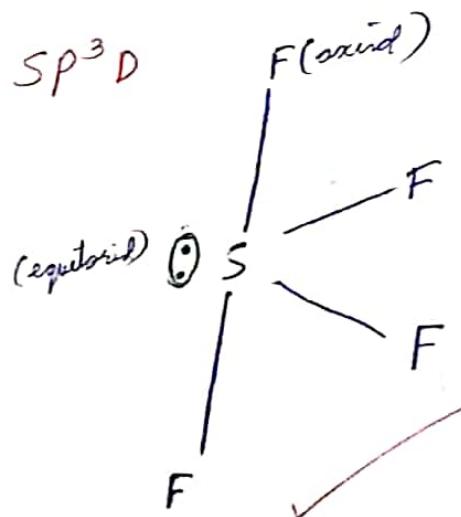
$$\sigma = 2$$

$$L.P = 2$$

$$S.N = 4$$

→ Bond angle of Ammonia & H_2O should be 109.5° according to hybridization 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.

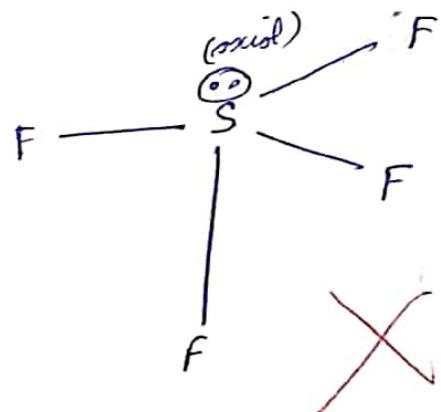
③ SP^3D



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

$$at 90^\circ \\ L.P - B.P = 2$$

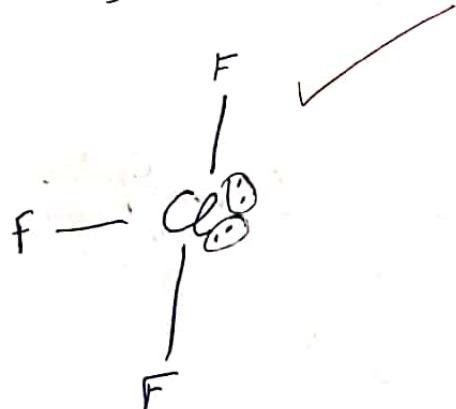
↓ less repulsion



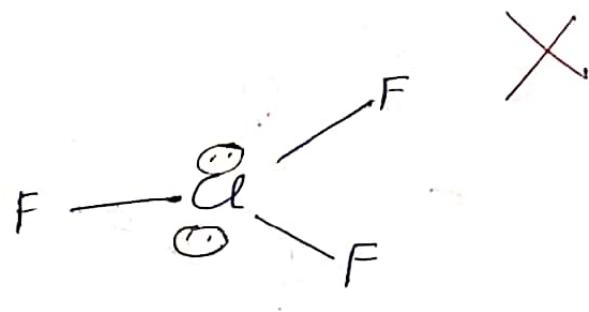
$$at 90^\circ$$

$$L.P - B.P = 3$$

ClF_3 ($\alpha = 3, \text{LP} = 2$)



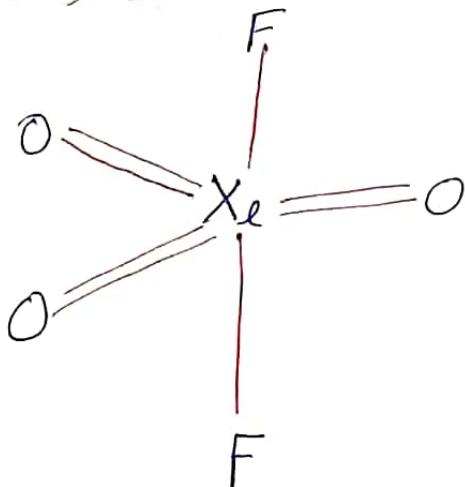
at 90°
 $\text{LP} - \text{BP} = 4$
less Repulsion



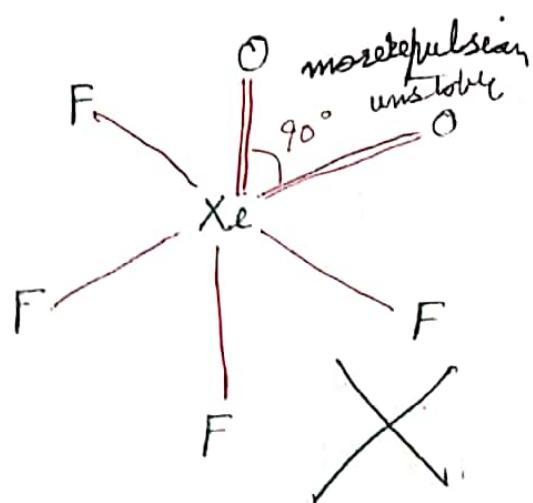
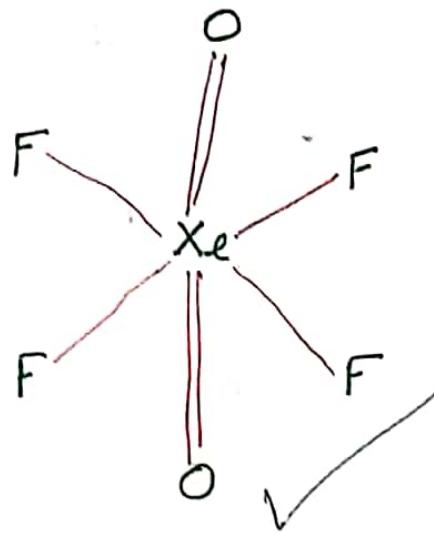
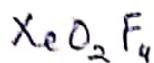
at 90°
 $\text{LP} - \text{BP} = 6$

→ Lone pair & Double bond will always be at
equatorial location.

~~XeO₃F₂~~



④ sp^3d^2 / Put Lone pair & Double bond at 180°

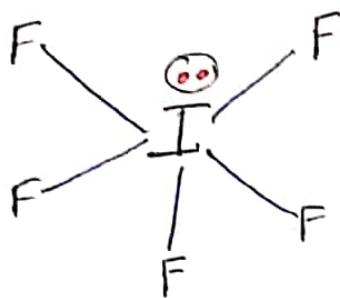


$$\sigma = 6$$

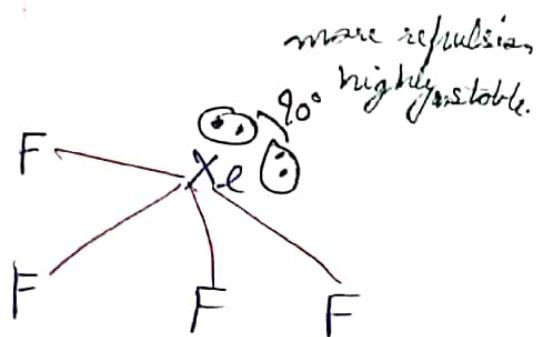
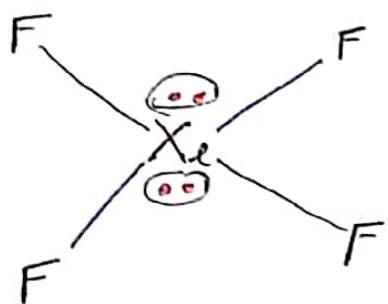
sp^3d^2

Square bipyramidal

Eg ① IF_5



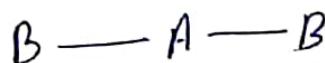
② XeF_4



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

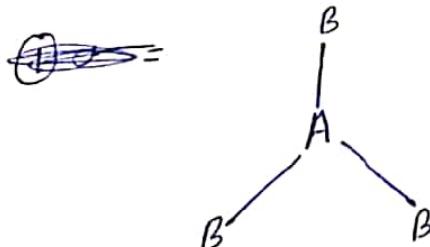
SP



- ① 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 = $180^\circ (\pm)$

- ⑧ Planar
- ⑨ max no. of atoms in one plane = 3

SP^2



- ① σ = 3
- ② L.P = 0
- ③ H.O = 3
- ④ SP^2
- ⑤ Trigonal Planar
- ⑥ Trigonal Planar
- ⑦ Bond angle = $120^\circ (3)$
- ⑧ Planar
- ⑨ 4

(63)

(3) AB_3E_{1-2} represent a lone pair
e.g. SiCl_2



(1) O = 2

(2) L.P = 1

(3) H.O = 3

(4) sp^2

(5) Trigonal Planar

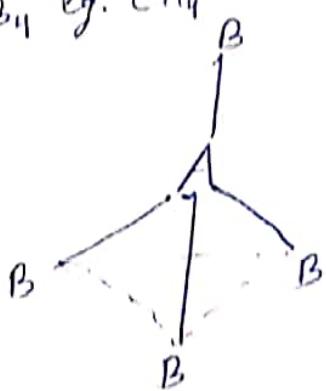
(6) V-shaped / Bent shape / Angular

(7) $<120^\circ$ (1)

(8) Planar

(9) 3

(4) AB_4 e.g. CH_4



(1) O = 4

(2) L.P = 0

(3) H.O = 4

(4) sp^3

(5) Tetrahedral

(6) Tetrahedral

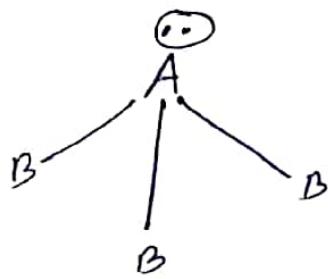
(7) 109.5° (6)

(8) Non-Planar

(9) 3

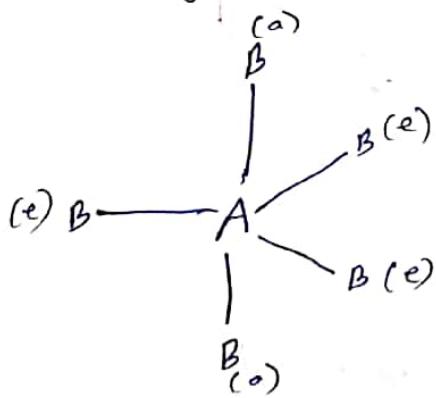
(64)

⑤ AB_3E_1 eg NH_3



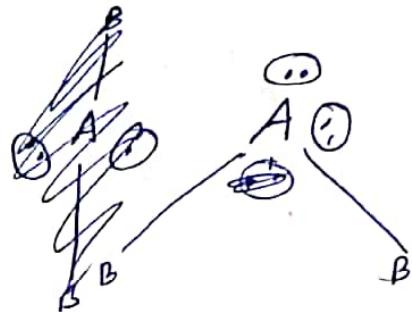
- ① $\sigma = 3$
- ② $L.P = 1$
- ③ $H.O = 4$
- ④ SP^3
- ⑤ Tetrahedral
- ⑥ Trigonal Pyramidal
- ⑦ ~~109.5°~~ $< 109.5^\circ (3)$
- ⑧ Non-Planar
- ⑨ ~~3~~ 3
 SP^3D

⑦ AB_5 eg PCl_5



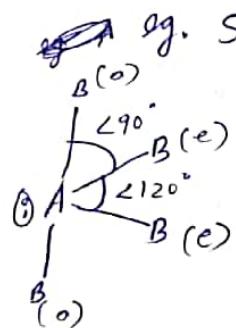
- ① $\sigma = 5$
- ② $L.P = 0$
- ③ $H.O = 5$
- ④ SP^3D
- ⑤ Trigonal Bipyramidal
- ⑥ Trigonal Bipyramidal
- ⑦ $120^\circ (3)$
- ⑧ $90^\circ (e)$
- ⑨ $180^\circ (1)$
- ⑩ Non-Planar
- ⑪ B^4

⑥ AB_2E_2 eg H_2O



- ① $\sigma = 2$
- ② $L.P = 2$
- ③ $H.O = 4$
- ④ SP^3
- ⑤ Tetrahedral
- ⑥ ~~V-shape~~
- ⑦ $< 109.5^\circ (1)$
- ⑧ ~~Plane~~
- ⑨ ~~3~~ 3

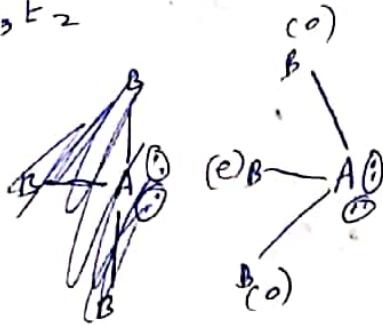
⑧ AB_4E_1 eg. SF_4



- ① $\sigma = 4$
- ② $L.P = 1$
- ③ $H.O = 5$
- ④ SP^3D
- ⑤ Trigonal Bipyramidal
- ⑥ ~~See-saw~~
- ⑦ $< 90^\circ (4)$
- ⑧ ~~$< 180^\circ (1)$~~
- ⑨ $< 120^\circ (1)$
- ⑩ Non-Planar
- ⑪ 3

⑨ $AB_3 E_2$

e.g. ClF_3



① $\sigma = 3$

② $L \cdot P = 2$

③ $H \cdot O = 5$

④ SP^3D

⑤ Trigonal BiPyramidal
Distorted T-Shaped

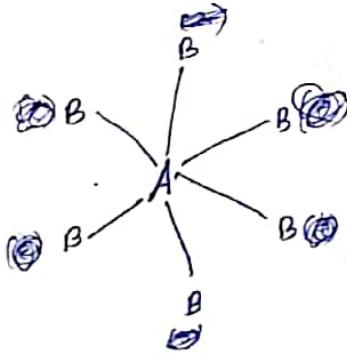
⑦ $< 90^\circ (2)$ $\leq 180^\circ (1) \alpha - \alpha$

⑧ Non-Planar

⑨ 4

SP^3D^2

⑩ $AB_5 E$ e.g. SF_5



① $\sigma = 5$

② $L \cdot P = 1$

③ $H \cdot O = 6$

④ SP^3D^2

⑤ Square BiPyramidal

⑥ Square BiPyramidal

⑦ $90^\circ (1)$

$180^\circ (2)$

⑧ Non-Planar

⑨ 5

⑪ $AB_2 E_3$ e.g. XeF_2



① $\sigma = 2$

② $L \cdot P = 3$

③ $H \cdot O = 5$

④ SP^3D

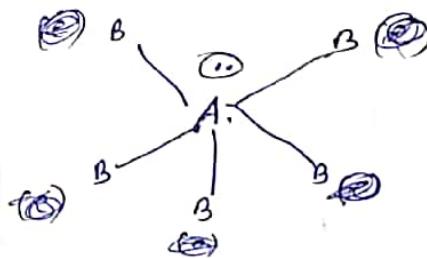
⑤ Trigonal BiPyramidal

⑦ $\approx 180^\circ (1) \alpha - \alpha$

⑧ Planar

⑨ 3

⑫ $AB_5 E_1$ e.g. $B_2 F_5$



① $\sigma = 5$

② $L \cdot P = 1$

③ $H \cdot O = 6$

④ SP^3D^2

⑤ Square BiPyramidal

⑥ Square Pyramidal

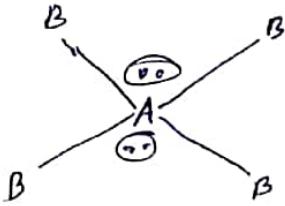
⑦ $< 90^\circ (3) - \text{around } 70^\circ$

$< 180^\circ (2) - \text{around } 175^\circ$

⑧ Non-Planar

⑨ 5

(13) AB_3E_2 eg XeF_4



$$\begin{array}{l} \textcircled{1} \quad \sigma = 4 \\ \textcircled{2} \quad L \cdot P = 2 \end{array}$$

$$\textcircled{3} \quad H \cdot O = 6$$

$$\textcircled{4} \quad SP^3D^2$$

~~(5)~~ Square Bipyramidal

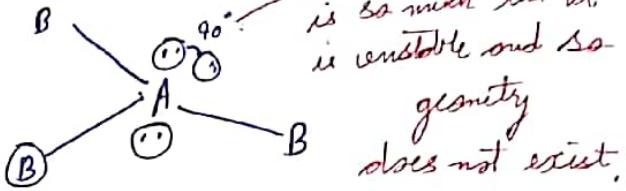
~~(6)~~ Square Planar

$$\textcircled{7} \quad 90^\circ(4) \quad 180^\circ(?)$$

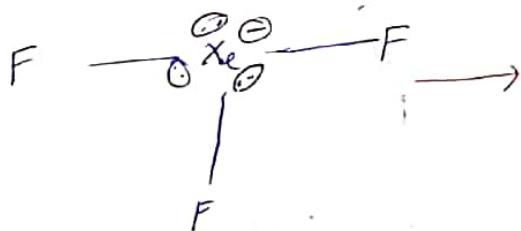
~~(8)~~ Planar

$$\textcircled{9} \quad 5$$

(14) AB_3E_3



Q XeF_3^- does not exist why



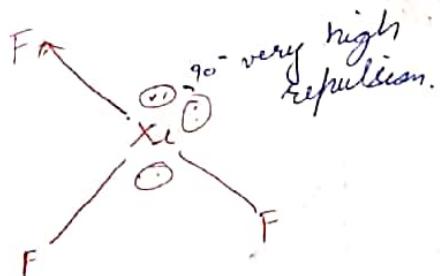
$$L \cdot P = 3$$

$$\sigma = 3$$

$$H \cdot O = 6$$

$$SP^3D^2$$

L-P-LP repulsion is so high
that it is unstable



H.W. 04-06-2024

O-1 (8-17)

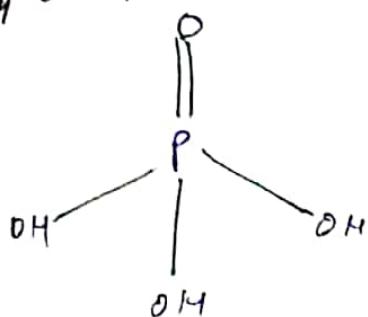
O-2 (6-11)

Roce - 13, 14

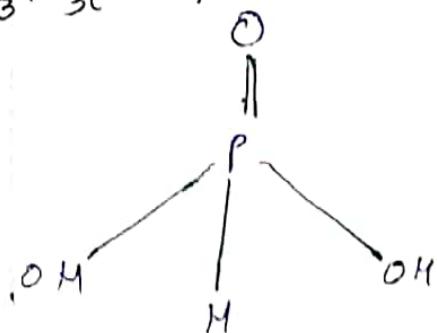
H.W. 06-06-2024

JM - 1, 2, 4, 5, 7, 8, 14, 16, 23, 24, 26, 27, 29, 31,
33, 42, 45, 46, 47, 53, 54, 57, 58

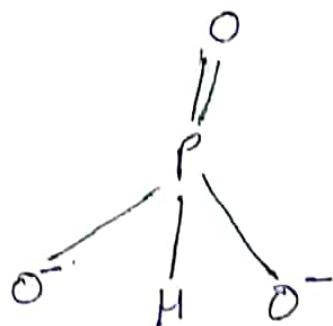
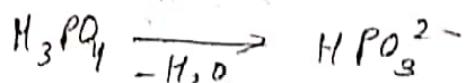
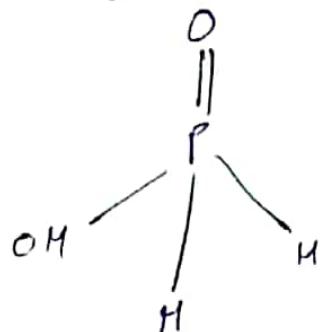
H_3PO_4 (Phosphoric acid)



H_3PO_3 (Phosphorous acid)



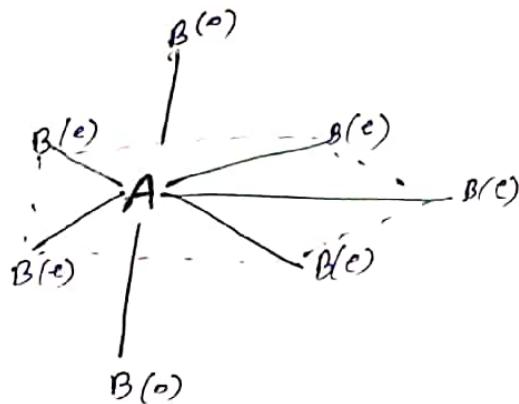
H_3PO_2 (Hydrophosphorous acid)



~~H₂O~~
Note - Whenever L.P-L.P repulsion is present at 90° in SP^3D^2 hybridisation, structure does not exist.

SP^3D^3

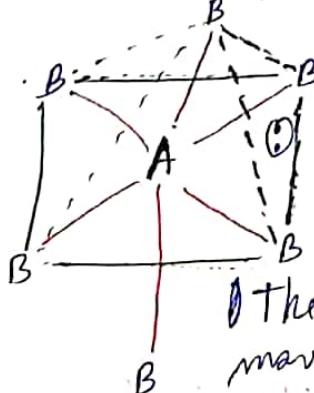
⑯ AB_7 eg IF_7



- ① $\sigma = 7$
- ② $L.P = 0$
- ③ $H.O = 7$
- ④ SP^3D^3
- ⑤ Pentagonal BiPyramidal
- ⑥ Pentagonal BiBipyramidal
- ⑦ $72^\circ(5) - (e-e)$
- ⑧ $90^\circ(10) - (o-o)$
- ⑨ $180^\circ(1) - (a-a)$
- ⑩ $144^\circ(5) - (e-o)$
- ⑪ Non-Planar
- ⑫ 6

~~Xod Raksha Hai~~

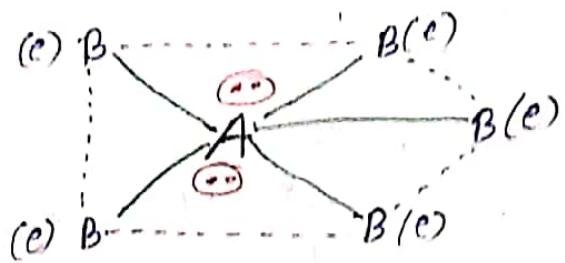
⑯ AAB_6E eg. $XeF_6, IF_5, XeOF_5^-$



The lone pair keeps moving in the block
From Feynman's note

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

(17) AB_5E_2 eg. XeF_5



① $d = 5$

② $L.P = 2$

③ $H.O = 7$

④ $SP^3D^3\delta$

⑤ Pentagonal Bipyramidal

⑥ ~~Planar~~ Bent and Planar

$SP^3D^3\delta$ & $L.P \rightarrow 2\pi$

180° & $\sqrt{5}\pi$

⑦ $72^\circ(s) - (c-c)$

$194^\circ(ss) - (c-c)$

⑧ Planar

⑨ 6

① ~~Draw~~ Draw Structure

① PCl_4

② IO_3^-

③ SO_2

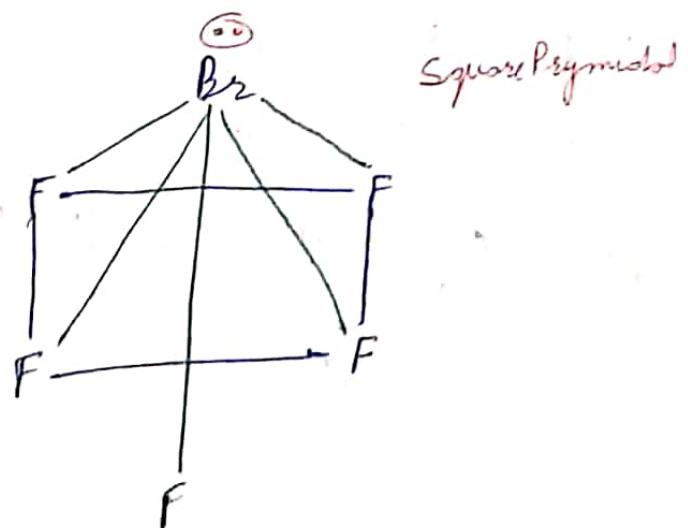
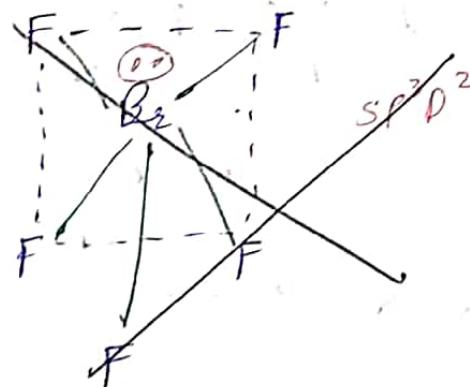
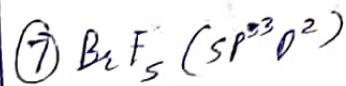
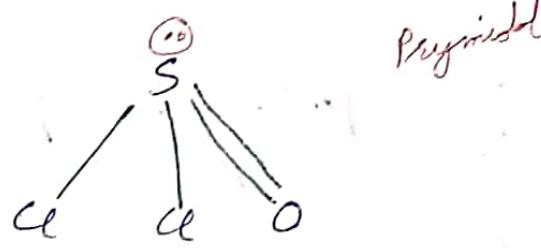
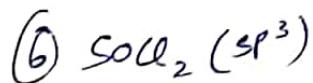
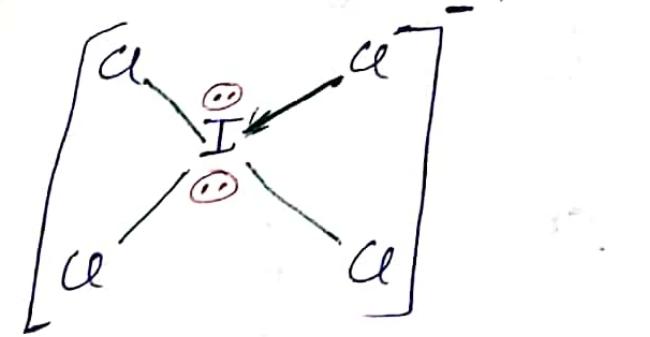
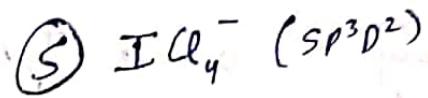
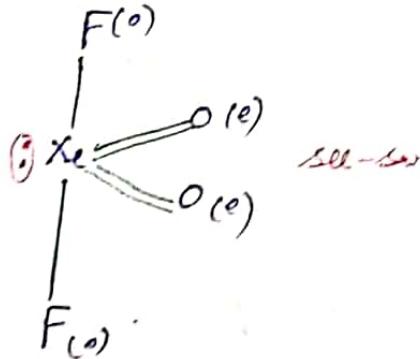
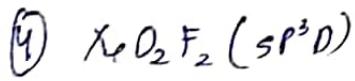
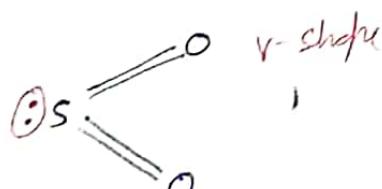
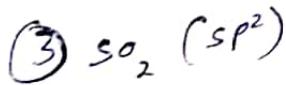
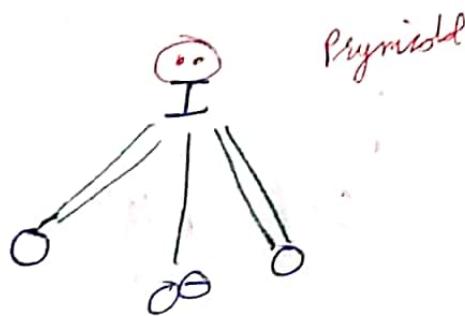
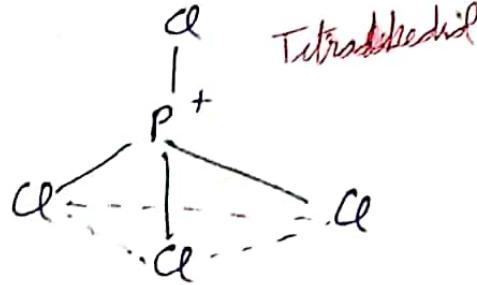
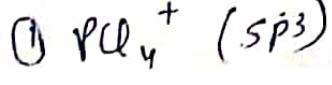
④ XeO_2F_2

⑤ ICl_4^-

⑥ $SOCl_2$

⑦ BzF_5

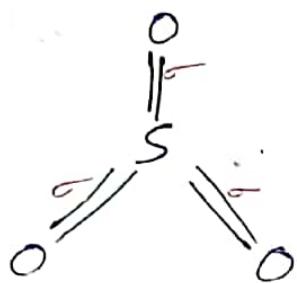
(70)



Q. SO_3 How many π -bonded types of π -bonds

Type of π -bond calculation

SP^2



S: $\begin{array}{|c|} \hline 1L \\ \hline \end{array}$ $\begin{array}{|c|c|c|} \hline 1L & 1 & 1 \\ \hline 3S & 3P & \\ \hline \end{array}$ $\begin{array}{|c|c|c|} \hline & 1 & 1 \\ \hline & 3d & \\ \hline \end{array}$

O: $\begin{array}{|c|} \hline 1L \\ \hline \end{array}$ $\begin{array}{|c|c|c|} \hline 1 & 1 & 1 \\ \hline 2S & 2P & \\ \hline \end{array}$

SP^2 : $\begin{array}{|c|} \hline 1L \\ \hline \end{array}$ $\begin{array}{|c|c|c|} \hline 1 & 1 & 1 \\ \hline 3S & 3P & \\ \hline \end{array}$ $\begin{array}{|c|c|c|} \hline 1 & 1 & 1 \\ \hline 3d & 3d & \\ \hline \end{array}$

SP^2 : $\begin{array}{|c|} \hline 1L \\ \hline \end{array}$ $\begin{array}{|c|c|c|} \hline 1 & 1 & 1 \\ \hline 3S & 3P & \\ \hline \end{array}$ $\begin{array}{|c|c|c|} \hline 1 & 1 & 1 \\ \hline 3d & 3d & \\ \hline \end{array}$

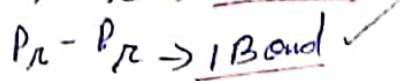
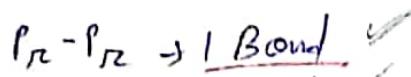
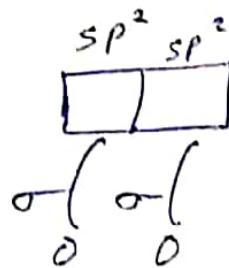
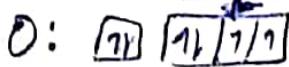
Diagram showing the formation of π -bonds from SP^2 hybrid orbitals. It shows two sets of O atoms with $2p$ orbitals. Red arrows indicate the movement of electrons from the $2p$ orbitals of one set to form π -bonds between the S atom and the O atoms.

$P_{\alpha} - P_{\alpha} \rightarrow 1 \text{ Bond}$
 $P_{\alpha} - P_{\beta} \rightarrow 2 \text{ Bond}$

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

Q1. SO_2 how many regions & types of regions.

① SO_2



② ClO_2^+

③ XeO_4^-

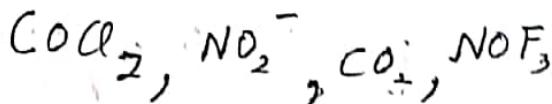
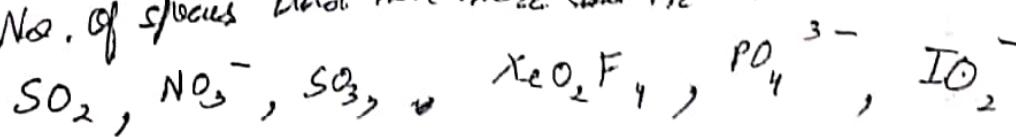
④ ClO_4^-

⑤ C_2H_2

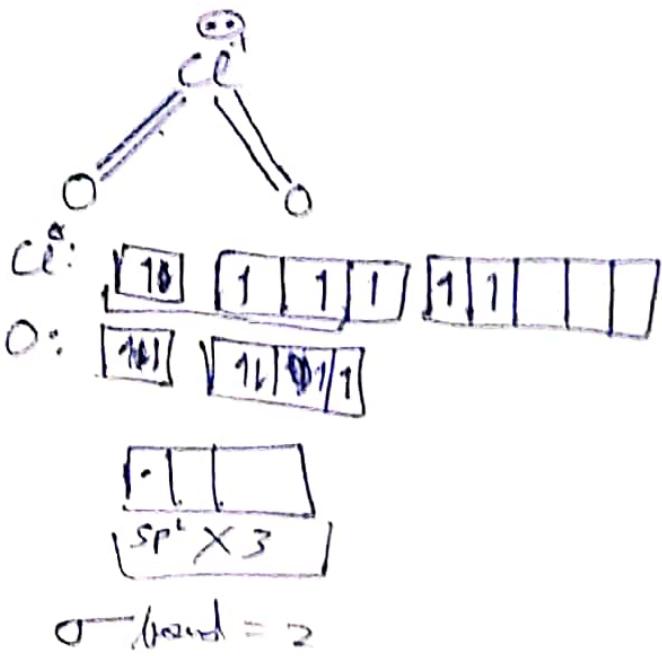
⑥ CO_3^{2-}

⑦ ClO_3^-

Q2. No. of species that have more than $\text{P}_R - \text{P}_R$ than $\text{P}_R - \text{P}_R$ bonds :-



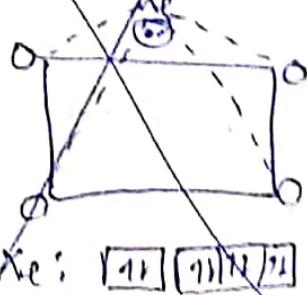
Q1 ② $\text{ClO}_3^- (\text{sp}^2)$



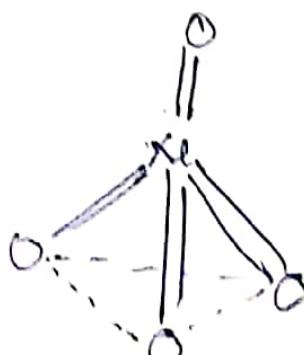
$$2P - 2P_N + 3P_N = 4 \quad \checkmark$$

$$-2P_N + 3D_N = 0 \quad \checkmark$$

③ $\text{XeO}_4^- (\text{sp}^3\delta^2)$

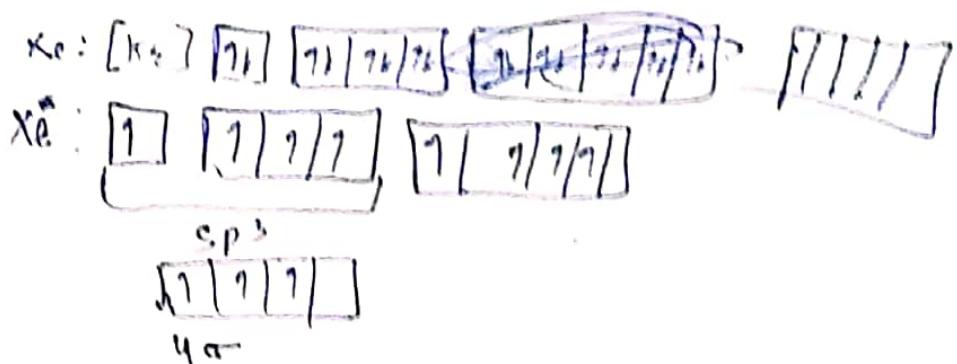


③ $\text{XeO}_4^- (\text{sp}^3)$

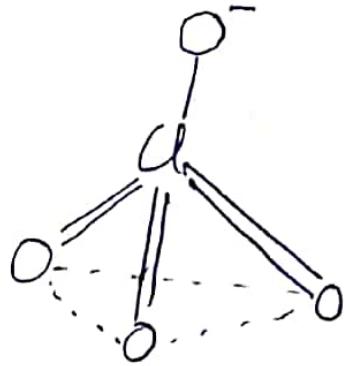


$$(2P_N - 5D_N) \quad P_N - P_N = 0 \quad \checkmark$$

$$P_N - P_N = 0$$



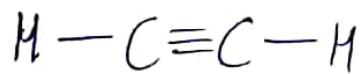
④ ClO_4^- (sp^3)



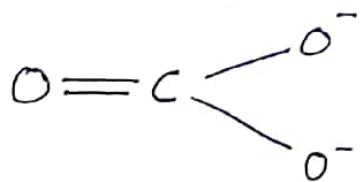
$$P_R - P_{R\bar{R}} = 0 \checkmark$$

$$P_R - D_R = 3 \checkmark$$

⑤ ~~C_2H_2~~ (sp)



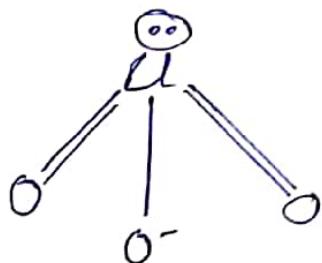
⑥ CO_3^{2-} (sp^2)



$$P_R - P_{R\bar{R}} = 1 \checkmark$$

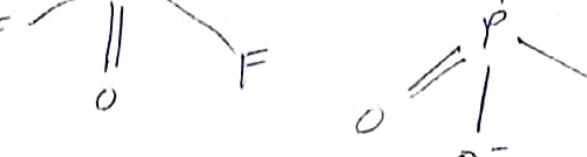
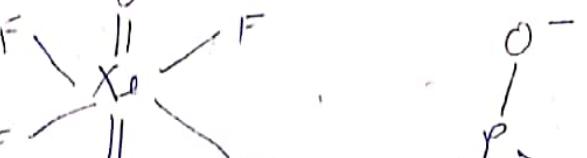
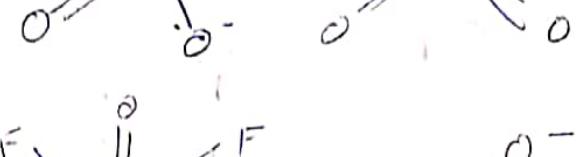
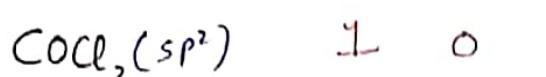
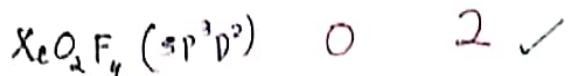
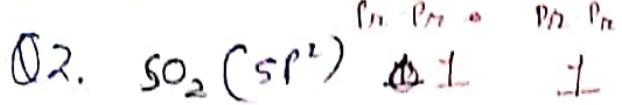
$$P_R - D_R = 0 \checkmark$$

⑦ ClO_3^- (sp^3)

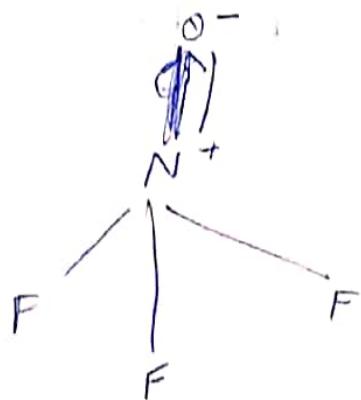
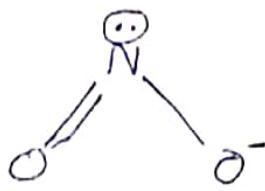
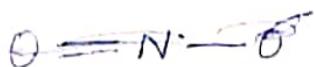
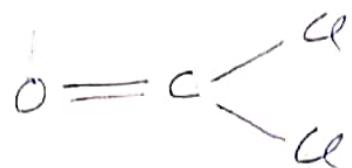


$$P_R - P_{R\bar{R}} = 0 \checkmark$$

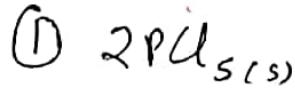
$$P_R - D_R = 2 \checkmark$$



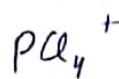
{ Q4 }



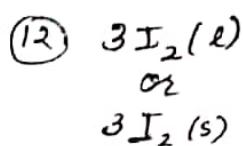
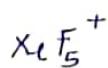
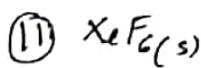
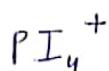
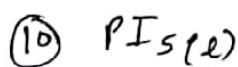
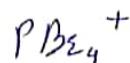
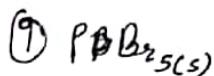
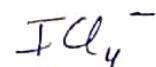
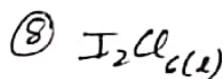
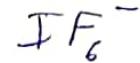
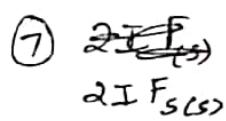
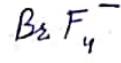
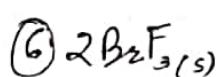
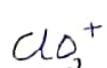
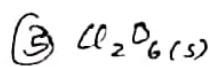
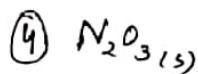
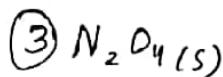
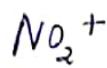
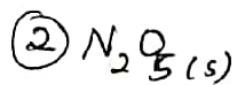
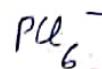
Solid/Liquid state hybridisation (Rotra Ha)



Cationic Part

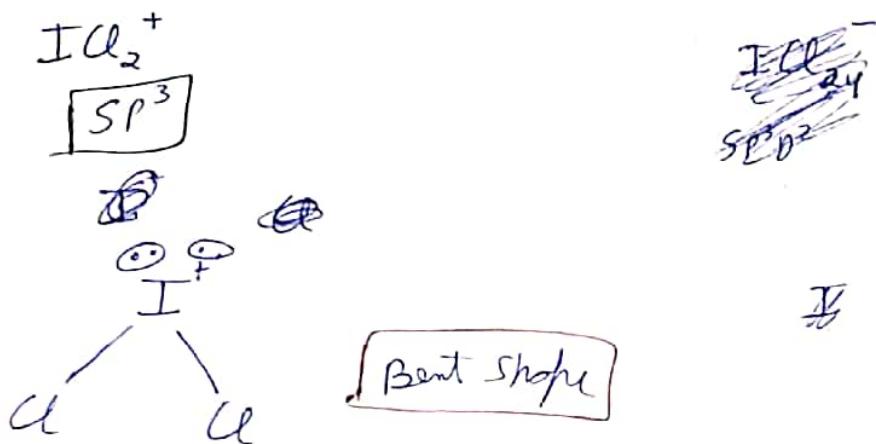


Anionic Part

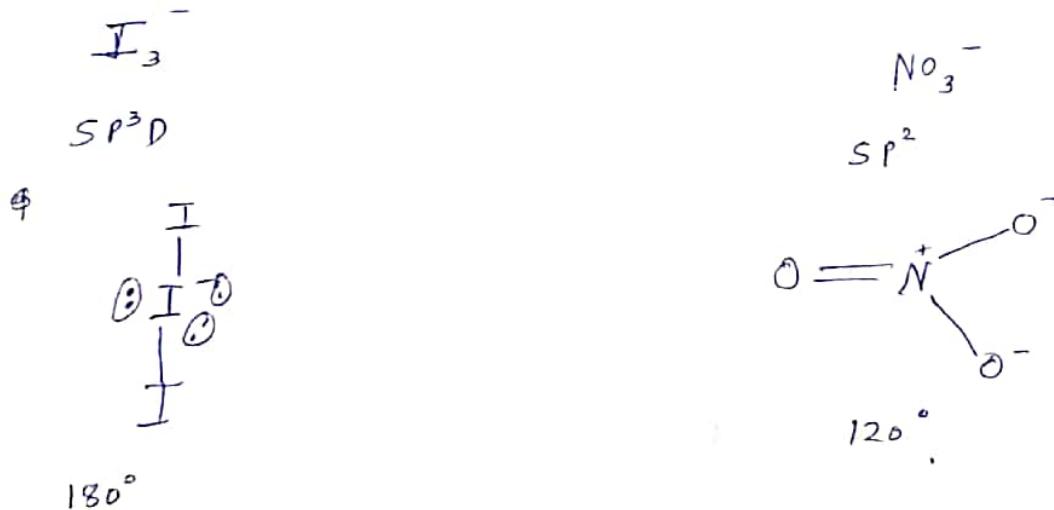


Q1

Q1 Comment on hybridisation & shape of cationic part of $I_2Cl_6^{(+)}$



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



$$180^\circ - 120^\circ$$

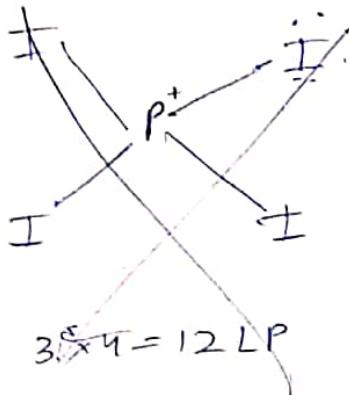
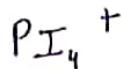
$$60^\circ$$

Q Comment hybridisation of anionic Part of $PBr_5(s)$

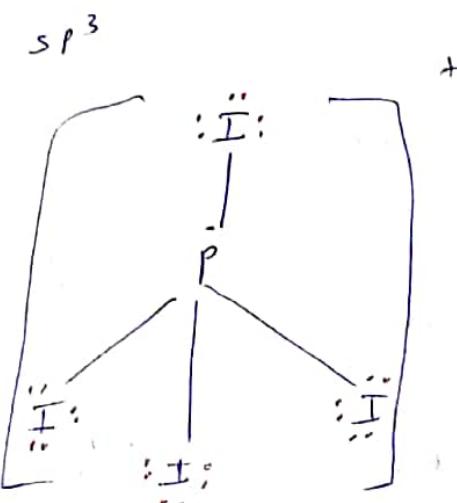


No hybridisation

Q Comment Total no. of L.P on cationic Part of $PI_5(s)$



$$3 \times 4 = 12 \text{ LP}$$



$$\boxed{L.P = 12}$$

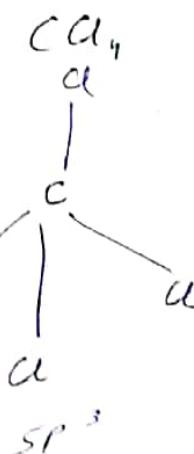
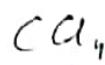
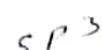
Bond Parameter :-

① Bond angle :-

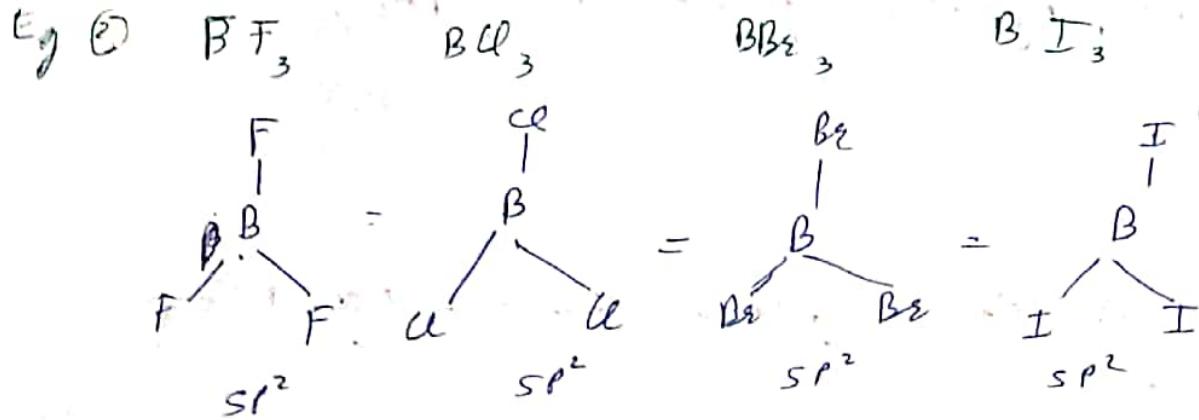
Rule 1 → If hybridisation is same & There is no Lone pair on central atom, Then angle remains same irrespective of size of atom. (surrounding atoms must be same some in a particular molecule)



$$L.P = 0$$



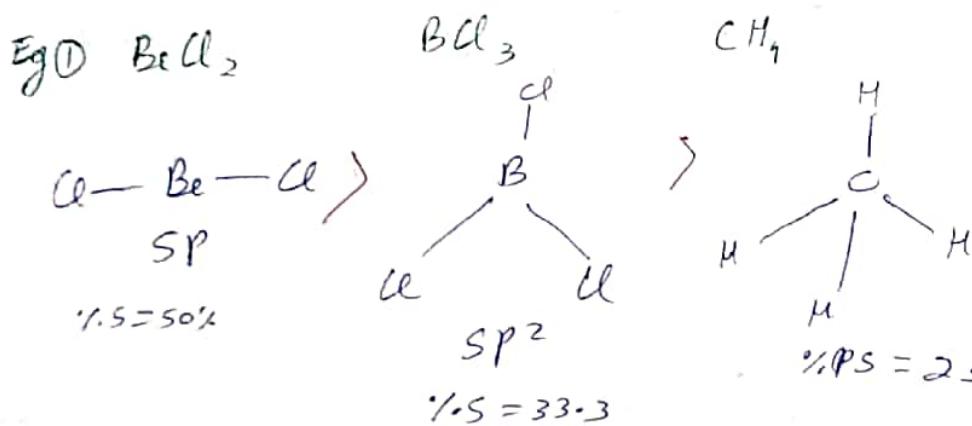
$$BA = \text{same} = 109.5^\circ$$



$$\angle \text{P} = 0$$

$$\text{Bond Angle} = \text{sum} = 120^\circ$$

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

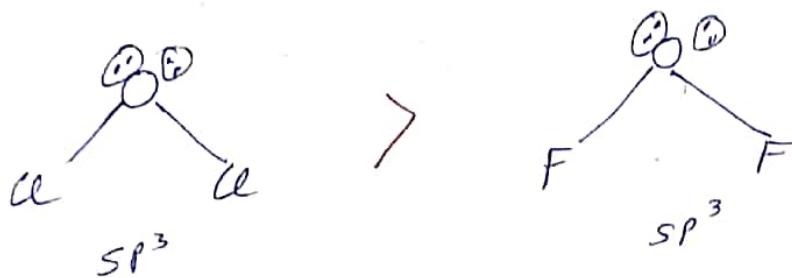


B. A & % S character

Rule 3:- If some type of hybridisation takes place, then check steric crowding or steric repulsion or steric hindrance.

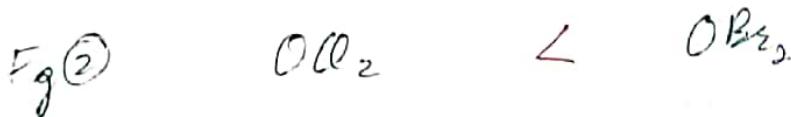
Size of surrounding atom ↑ B·A↑

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 (C₁₁H₂₄, C₂H₅ etc) Then steric crowding will appear.



size of Cl > size of F

$$\text{BA}(\text{OCl}_2) > \text{BA}(\text{OF}_2)$$



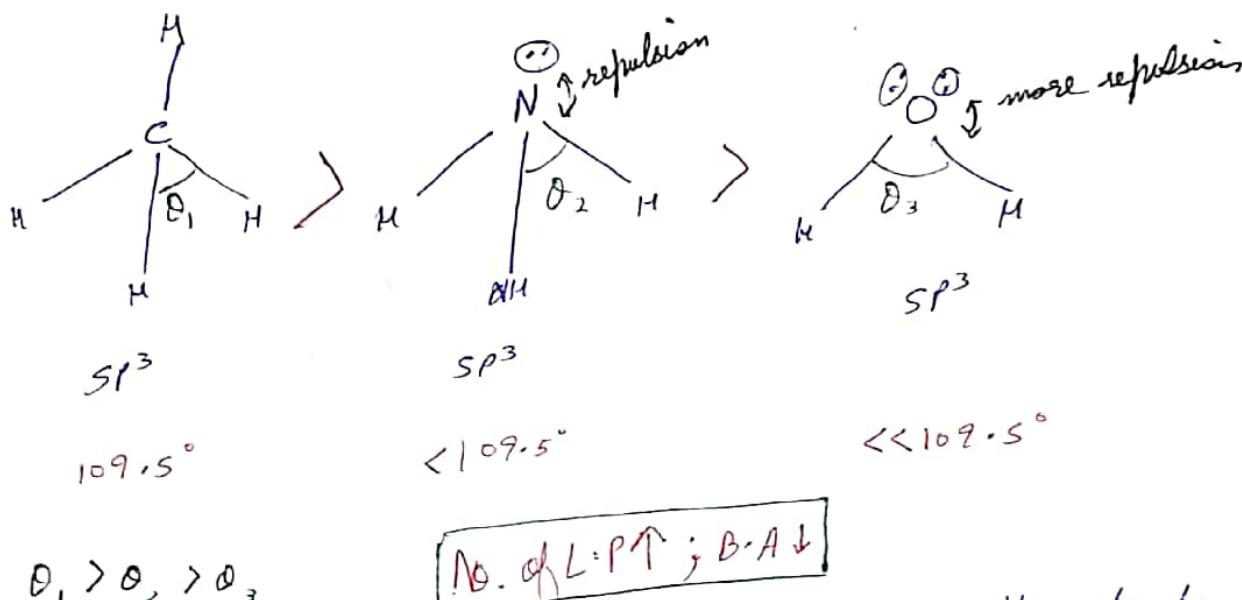
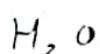
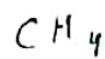
size of Cl & size of Br

$$\text{BA(Cl)} < \text{BA(Br)}$$

Rule 4:- If some 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.

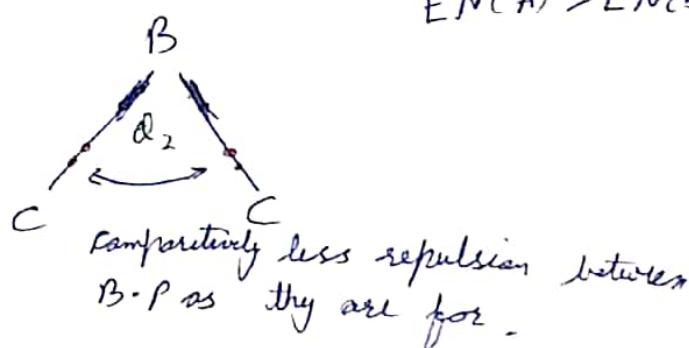
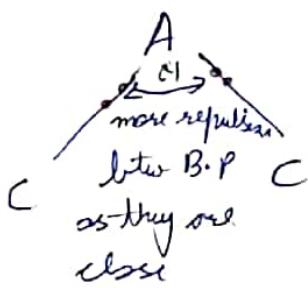
Eg ①



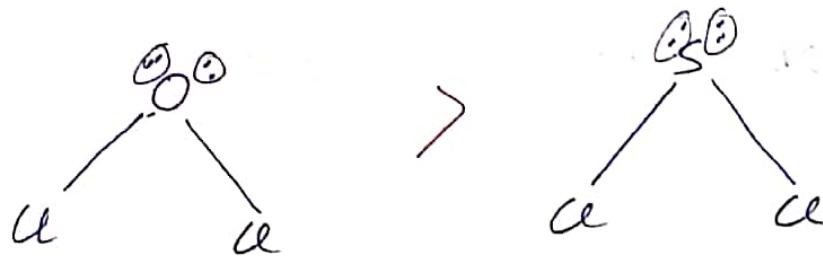
Rule 5:- If all above factors are not working then check electron negativity of central atom or surrounding atom.

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

$$\text{EN}(A) > \text{EN}(B)$$

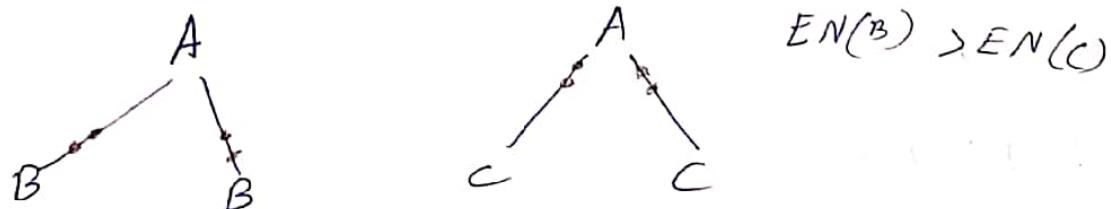


Eg ①. OCl_2



EN of central atom ↑ ; B-A ↑

ii) EN of surrounding atom {Same central atom}



EN of surrounding atom ↑ ; B-A ↓

Eg. OH_2 > OF_2

$\text{EN}(\text{H}) < \text{EN}(\text{F})$

$\text{BA.}(\text{OH}_2) > \cancel{\text{BA.}}(\text{OF}_2)$

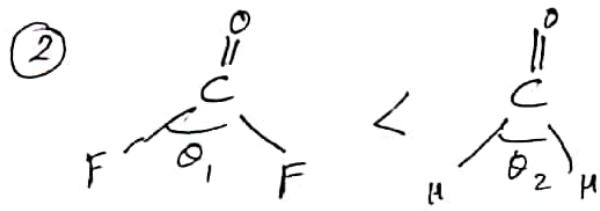
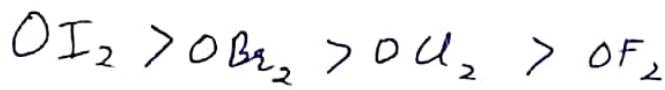
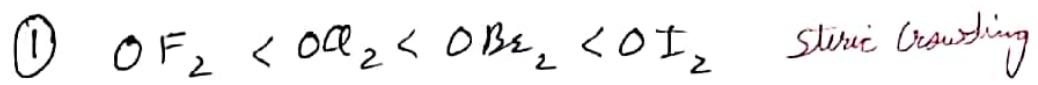
Rule 6 - Multiple bond repel more strongly as compared to single bond.

$(\text{PB} - \text{P.B}) > (\text{PB-SB}) > (\text{SB-SB})$

Eg ① COH_2



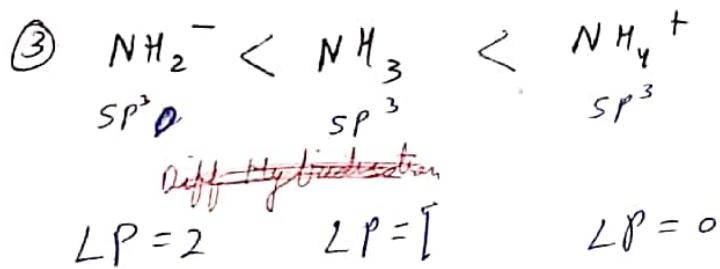
Q Bond angle order?



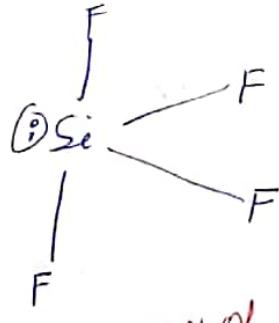
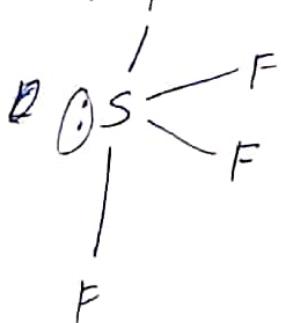
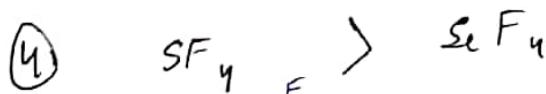
$$\theta_1 < \theta_2$$

EN of surrounding atom

$$EN(F) > EN(H)$$

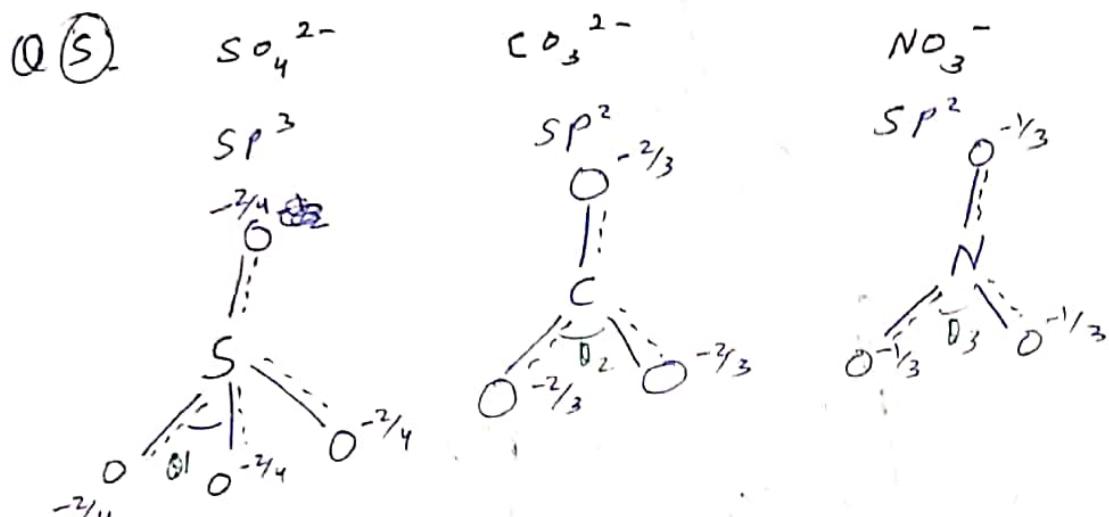
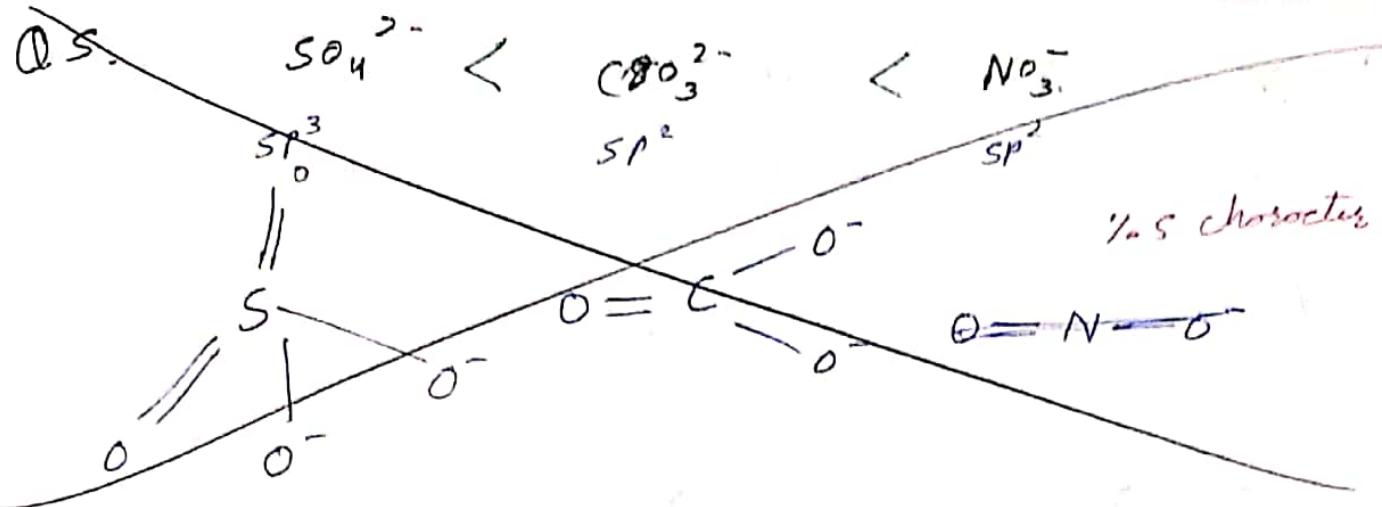


• \bullet Lone pair



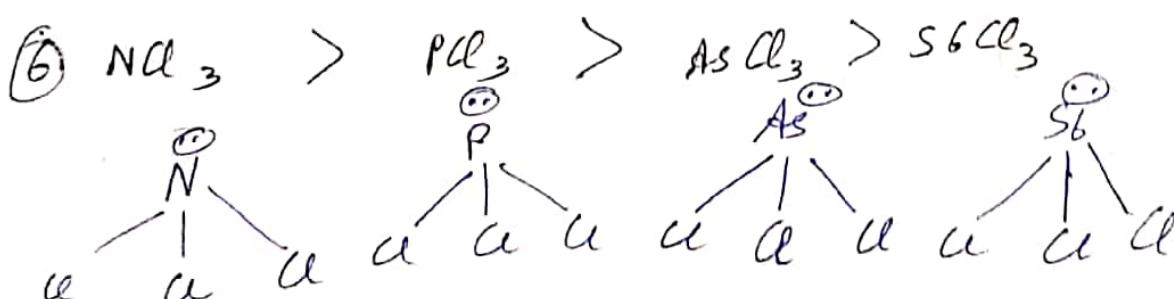
$$EN(S) > EN(Se)$$

EN of central atom

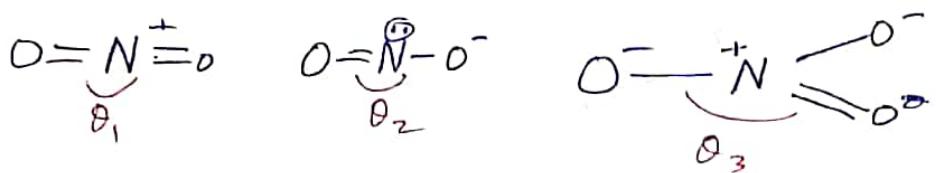
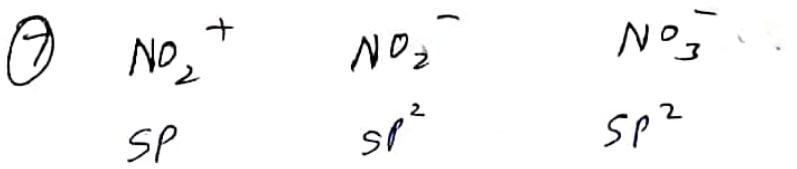


$$O_1 < O_2 = O_3 \quad \{ \text{B.A. \& V.S character} \}$$

~~By V.S character~~

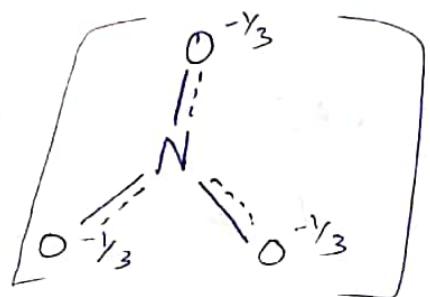


EN of central atom.

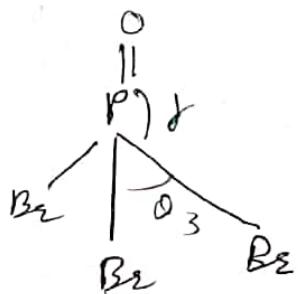
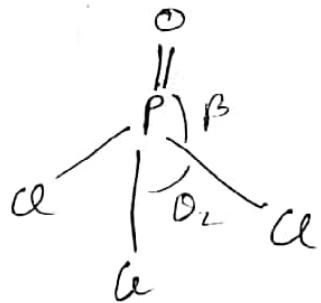
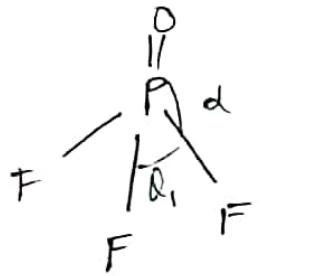


$$\boxed{\theta_1 > \theta_3 > \theta_2}$$

↓

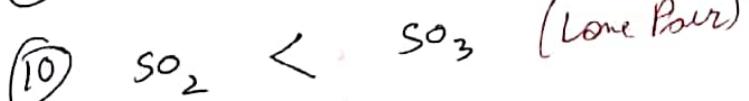
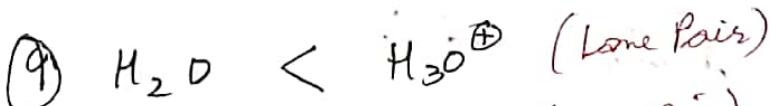


⑧



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

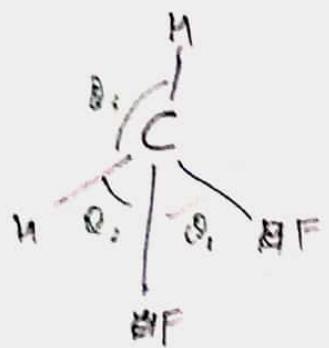
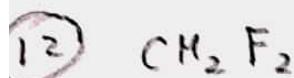
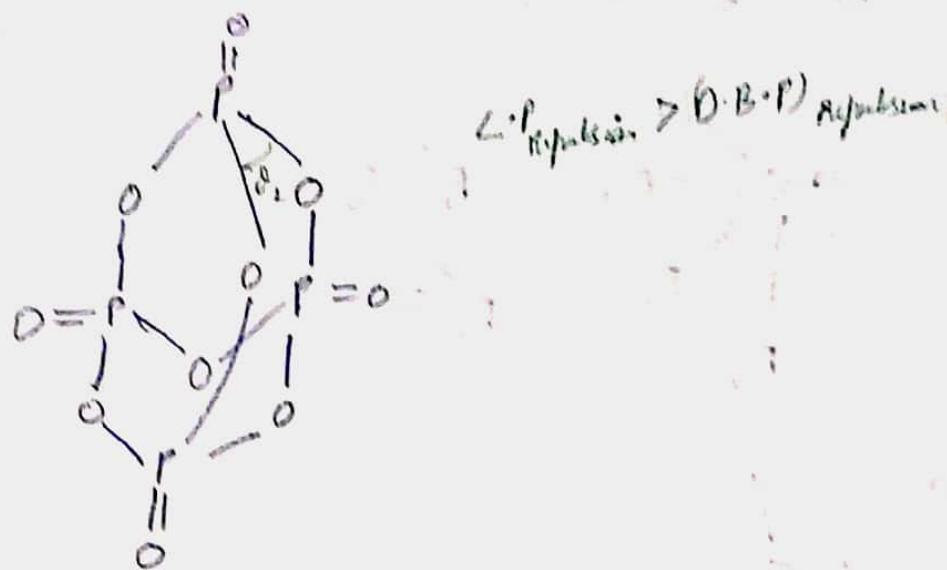
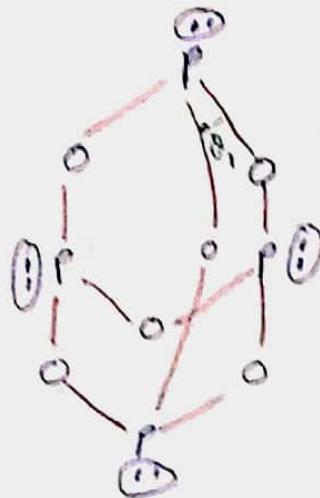
$\alpha > \beta > \gamma$ {depends on θ}



⑪

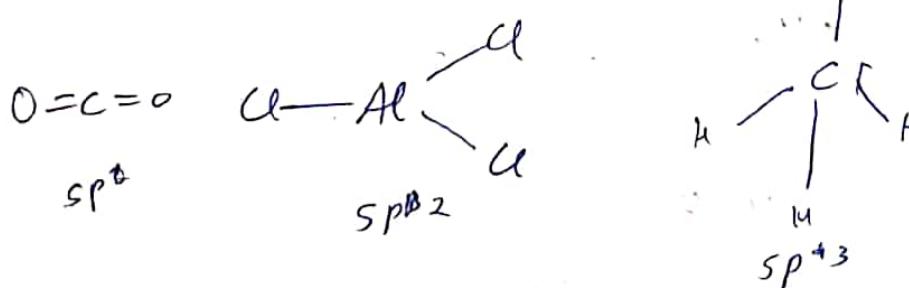
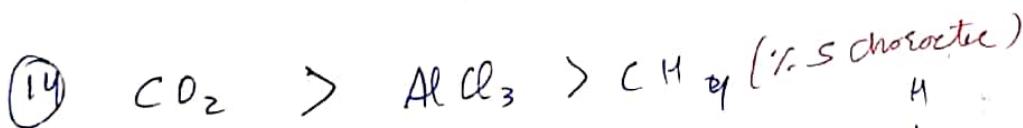
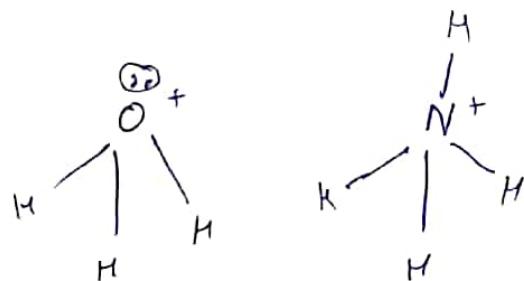
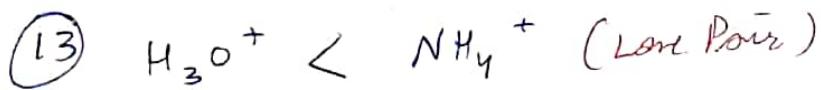
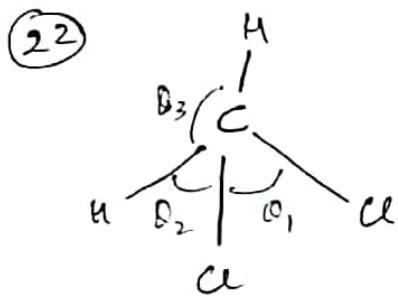
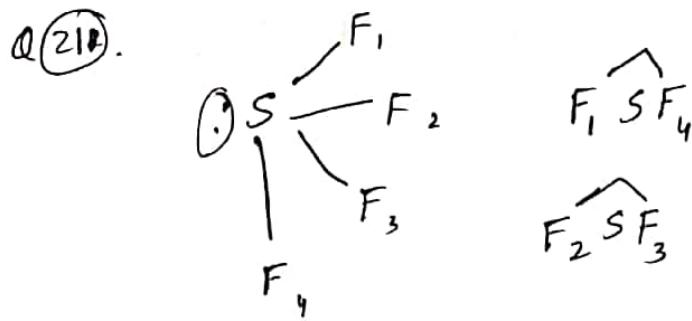
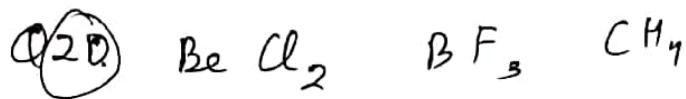


(86)

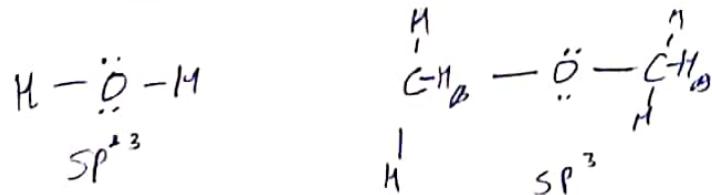
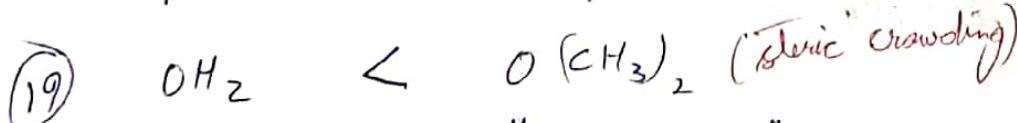
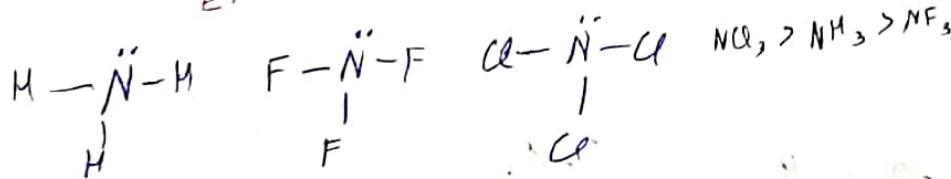
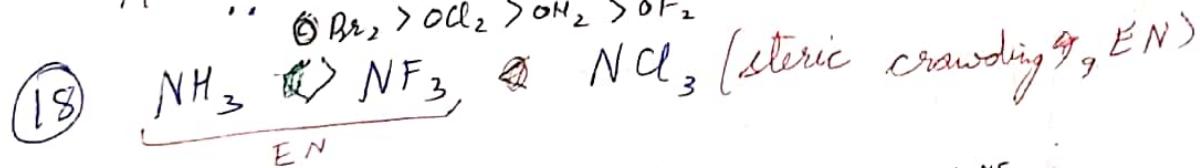
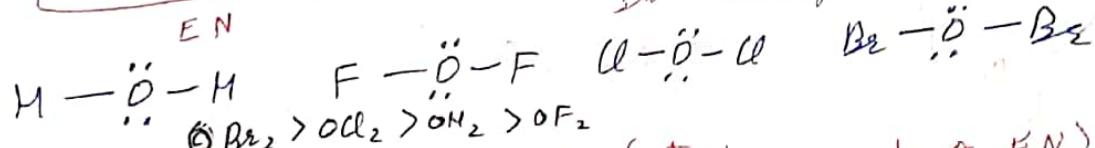
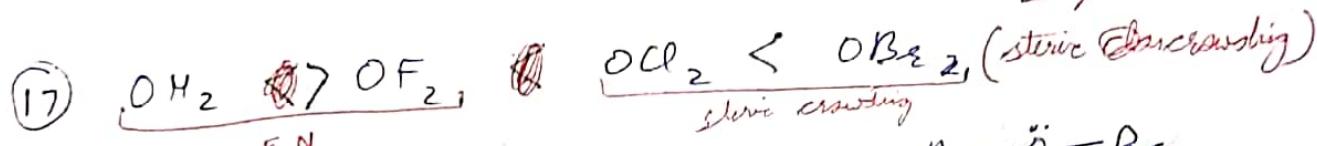
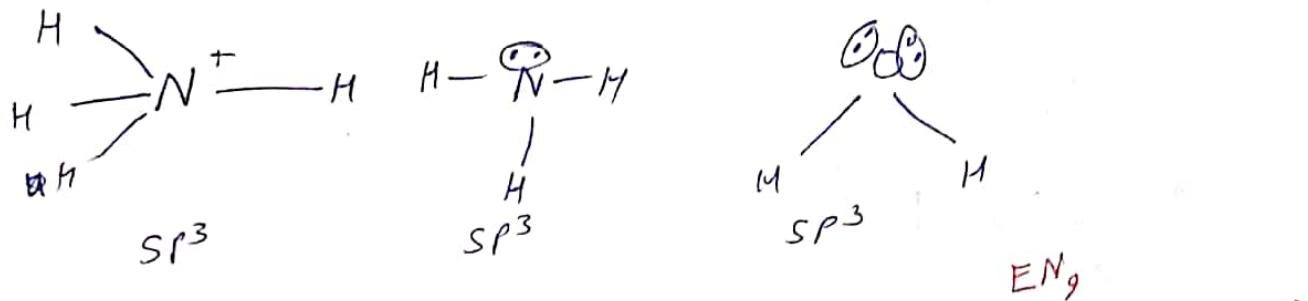
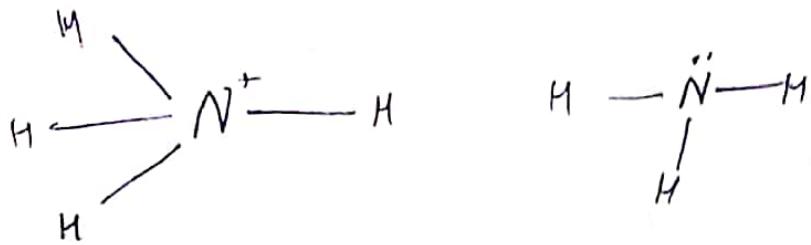


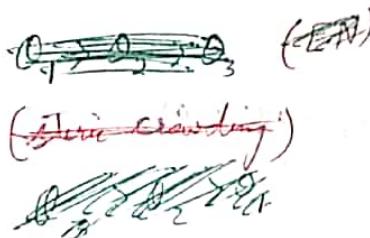
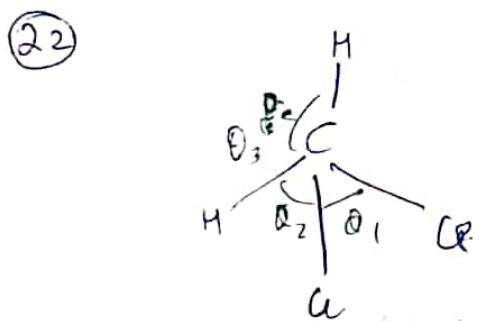
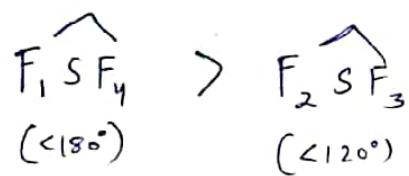
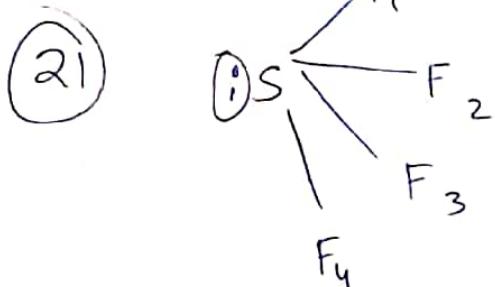
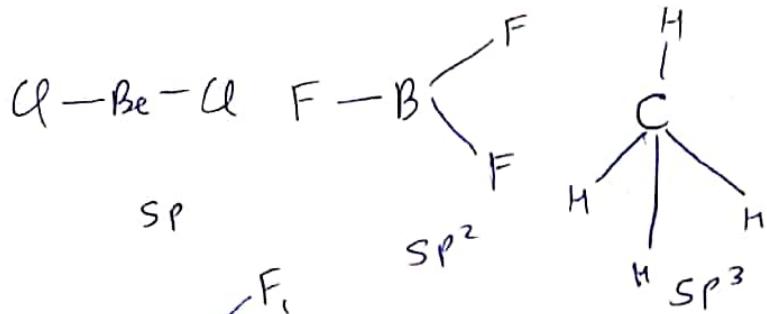
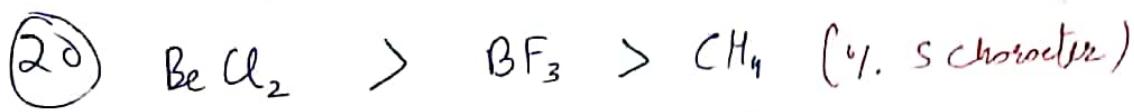
$O_3 > O_2 > O_1$ (EN, repulsion)



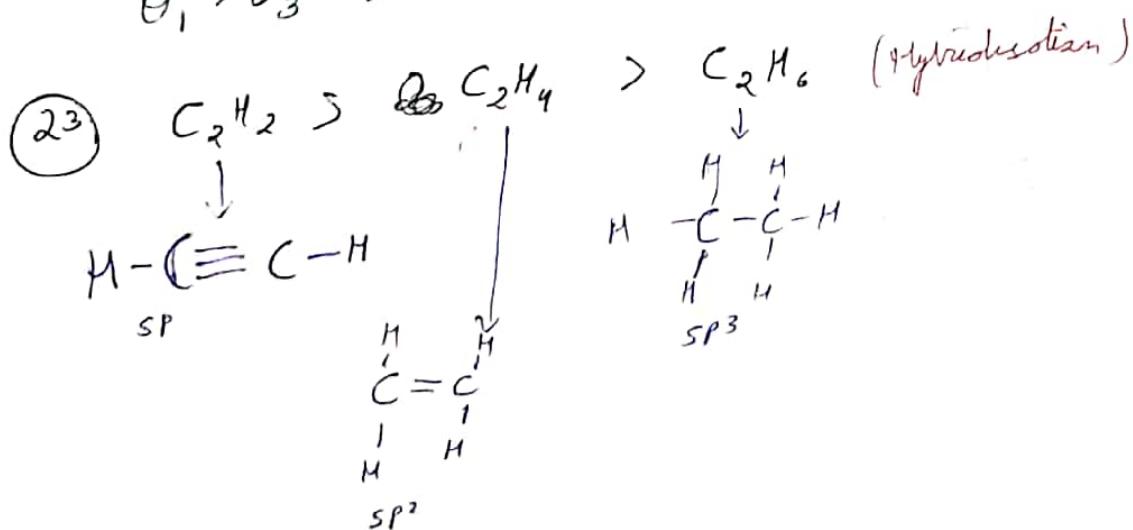


(88)



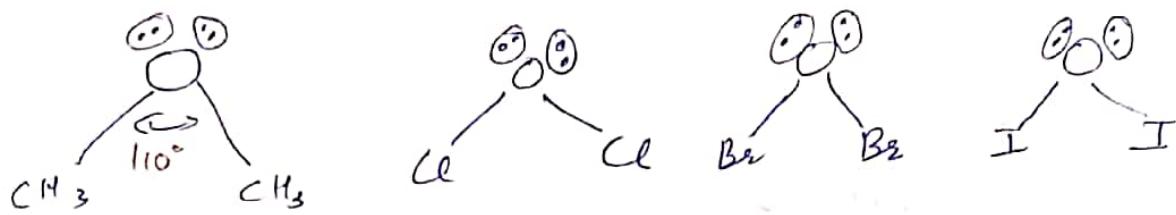


$$\theta_1 > \theta_3 > \theta_2$$



(90)

- (24) $\text{NF}_3 < \text{NCl}_3 < \text{NBr}_3 < \text{NI}_3$ (steric crowding)
- (25) $\text{O}(\text{CH}_3)_2 < \text{OCl}_2 < \text{OBr}_2 < \text{OI}_2$ (Exception)



Dobro's Rule :-

hybridization fail

B.A. $\text{NH}_3 > \text{PH}_3 > \text{AsH}_3 > \text{SbH}_3$

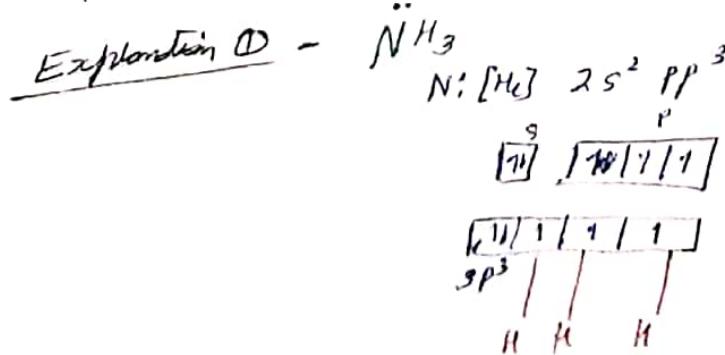
$\text{H}_2\text{O} > \text{H}_2\text{S} > \text{H}_2\text{Se} > \text{H}_2\text{Te}$

$\text{PH}_3 \rightarrow 94^\circ$
 $\text{AsH}_3 \rightarrow 92.5^\circ$
 $\text{SbH}_3 \rightarrow 91^\circ$

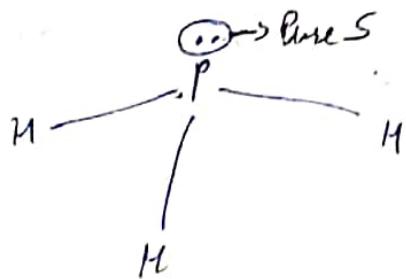
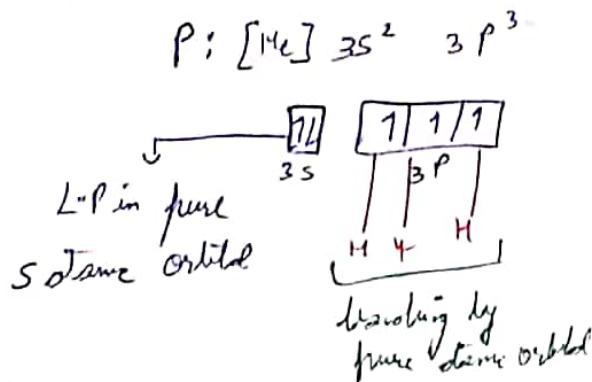
$\text{H}_2\text{S} \rightarrow 92^\circ$
 $\text{H}_2\text{Se} \rightarrow 91^\circ$
 $\text{H}_2\text{Te} \rightarrow 90^\circ - 5^\circ$

Angles are around 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 ≤ 2.01 . 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-orbital orbitals.



PH_3



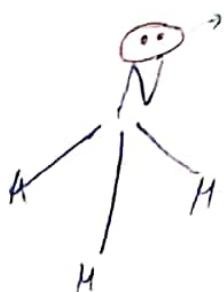
Explanation - 2 :- The cause of no hybridization in this molecule can be explained as follows.

→ In case of $\text{P}_3\text{P}_5\text{O}_10$, the energy required in hybridization is about 600 kJ/mol which is not compensated by energy released from the bond formation using hybrid orbitals.

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

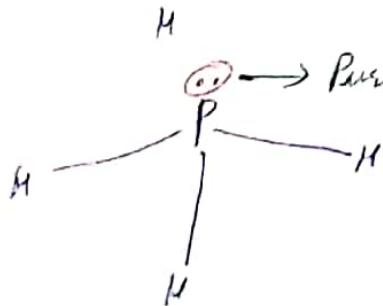
Q2. Formation of ammonium ion is much easier than phosphonium ~~ion~~ ion (PH_4^+) why?

Q3. Ammonia is easily soluble in water than PH_3 why?



$\rightarrow \text{sp}^3 \text{ orbital} \Rightarrow 25\% \text{ s character} \Rightarrow$ Attraction from nucleus is less

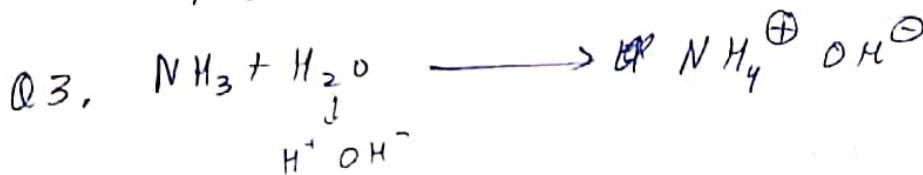
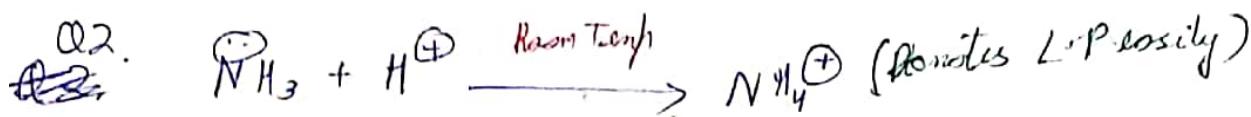
\downarrow
 e^- remove easily



$\rightarrow \text{Pure s orbital} \Rightarrow 100\% \text{ s character} \Rightarrow$ Attraction of nucleus is more

\downarrow
removing e^- requires more energy.

Q1. A Lewis base ~~to~~ donates lone pair easily & as $\text{NH}_3 \cdot \text{L.P}$ require very less energy to remove lone pair than PH_3 . Thus, NH_3 is stronger Lewis base.



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

→ In PH_3 (Phosphine), L.P of Phosphorus is present in ~~one~~ almost pure S-orbital which is less directional or less reactive.

④ Types of Hybridisation.

Equivalent Hybridisation

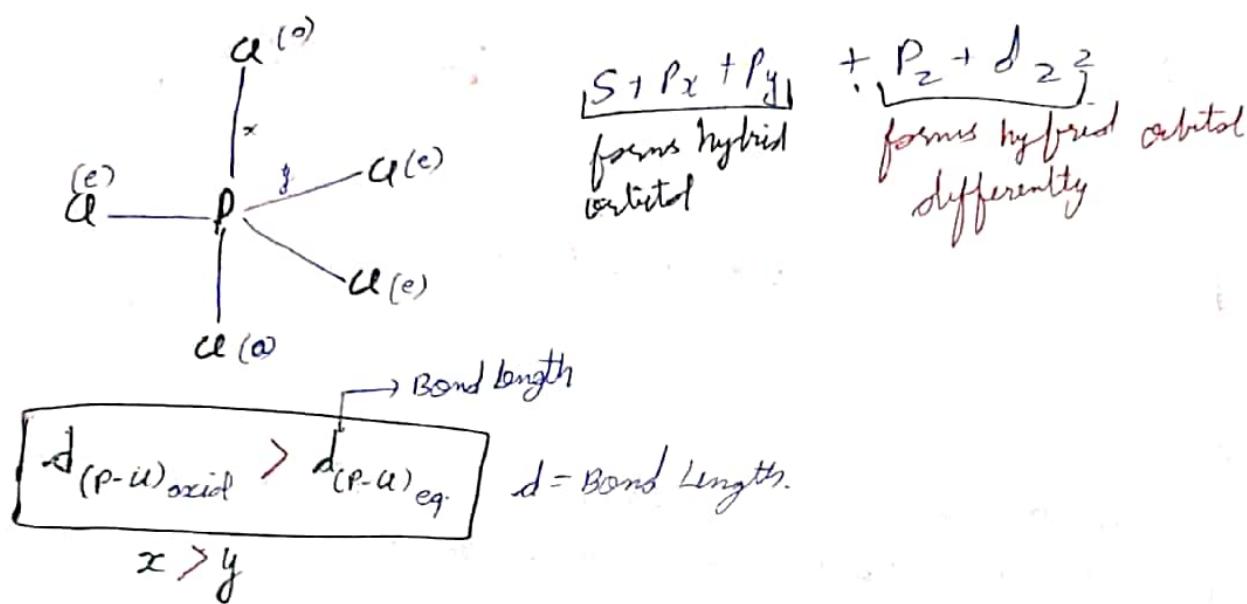
- no L.P on central atom
 - surrounding atoms must be same
 - All bond angles are same
- Eg $\text{CH}_4, \text{CCl}_4, \text{BF}_3$

Non Equivalent Hybridisation

- If there is L.P on central atom, surrounding atoms are not same or different bond angles
- Eg $\text{NH}_3, \text{CO}_2, \text{PCl}_5, \text{IF}_7, \text{XeF}_4$

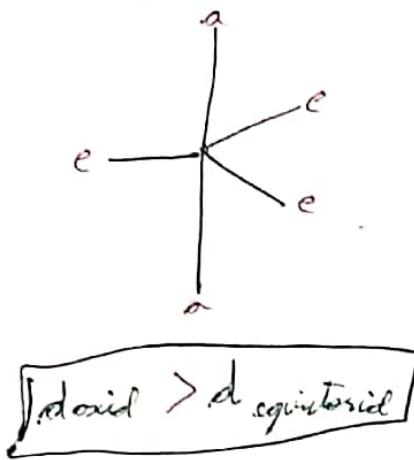
Bent's Rule :-

→ It is applicable for non-equivalent hybridisation.



- In $Pd_5(g)$ It is observed that there are two longer (weaker) axial bands and 3 shorter (stronger) equatorial bands are present.
- Experimentally It is concluded that all hybrid orbitals are not equal in Pd_5 .
- Equatorial bands are formed by sp^2 hybrid orbitals and axial bands are formed by pd hybrid orbitals of central atom.
- According to Bent's rule.
 - More Electron Negative element (surrounding atom) prefer to stay in the hybrid orbital of central atom. which is having less % s character 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 nucleus.

In SP^3D hybridisation (TBP)



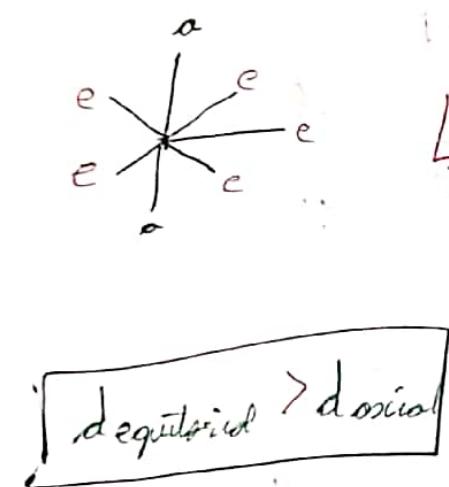
$$S + P_x + P_y + P_z + d_z^2$$

SP^2 PD

↳ directional nature ↑; $P \uparrow$, \downarrow
 S ↓, more EN will
 length of H-O ↑

↳ directional nature ↓; $\%P \downarrow$ length of H-O ↓
 $\%S \uparrow$ L-P or double will stay.

In SP^3D^3 hybridisation (PBP)



$$S + P_{3c} + P_y + D_{3c}y + D_{3c}^2 - y^2 + P_z + PD_z^2$$

SP^2P^2 PD

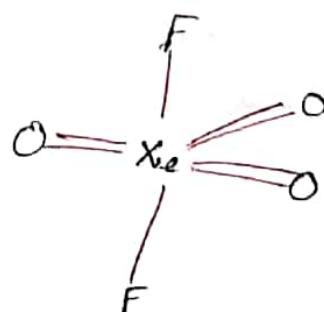
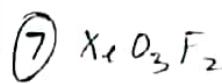
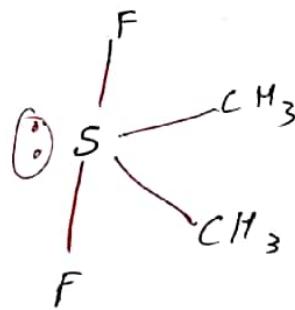
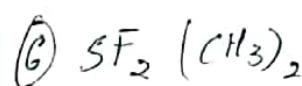
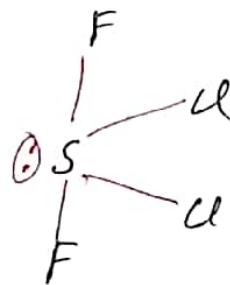
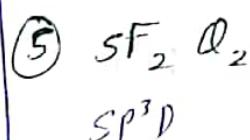
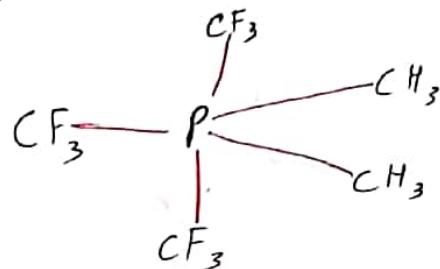
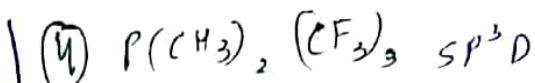
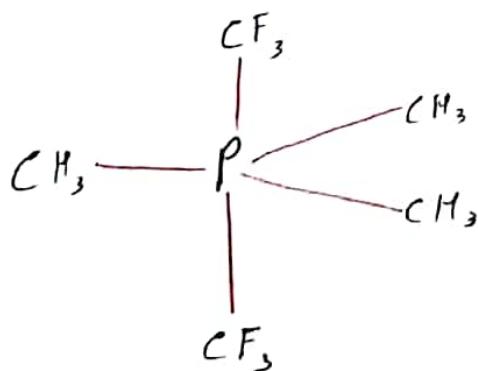
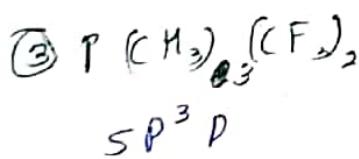
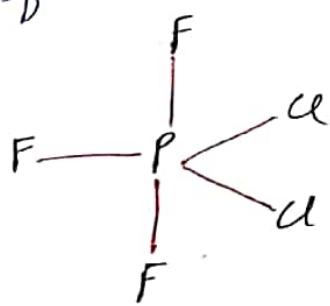
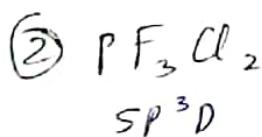
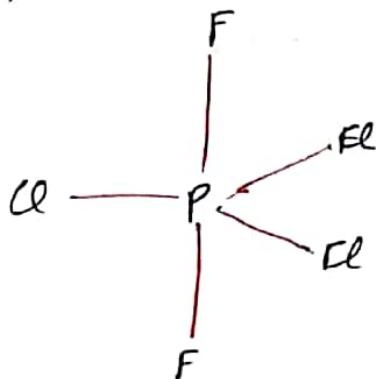
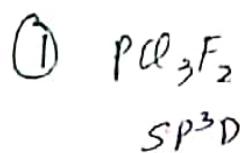
↳ directional nature ↑;
 $\%P \uparrow$; length of H-O ↑
 $\%S \downarrow$; more EN will stay

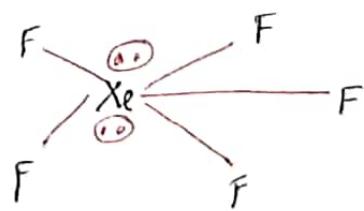
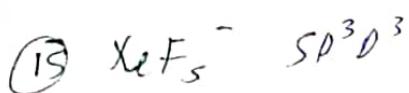
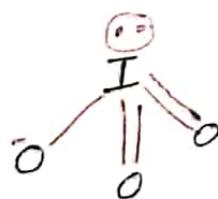
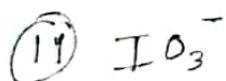
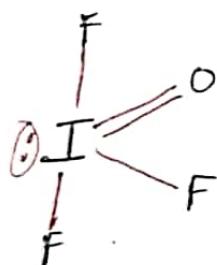
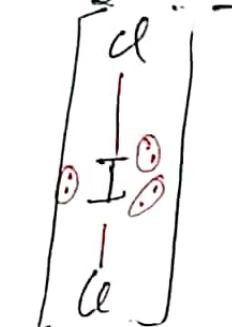
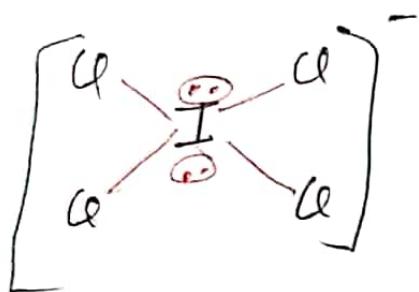
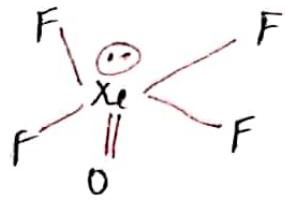
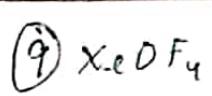
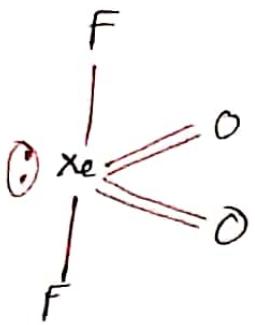
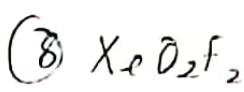
↳ directional nature ↓;
 $\%P \downarrow$; length of H-O ↓
 $\%S \uparrow$; ~~L-P~~ Double bond
 will stay.

Note:-

1. SP^3D - A strong \ominus electron Negative element ~~will~~ will prefer axial position, in SP^3D hybridisation.
2. A L-P or a double bond occupy equatorial ~~back~~ position in SP^3D^3 hybridisation because they want to be close to the nucleus.

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

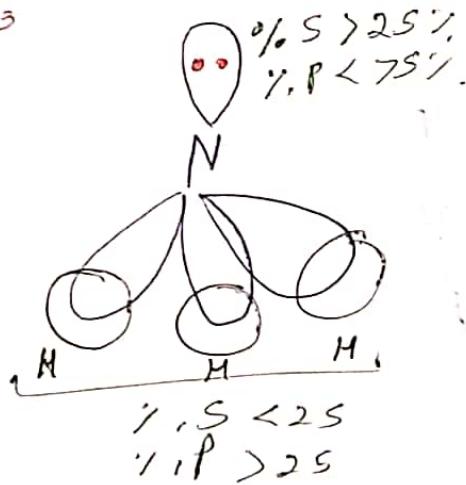




⑯

Alternative Statements of Bent's 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~~, which depends upon circumstances.



GET L-P ~~जटिल~~, %S character
GST ~~जटि~~

GET E N ~~जटिल~~ %S character
पर्याप्त जटि

How To solve Questions of Bond Length.

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

Shell no. ↑ Bond Length ↑

e.g. $H-F < H-Cl < H-Br < H-I$
 $d = (R_H + R_F) & (R_H + R_{Cl}) & (R_H + R_{Br}) & (R_H + R_I)$

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

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 some period.

left \rightarrow Zeff \uparrow Size \downarrow
right Bond length \downarrow

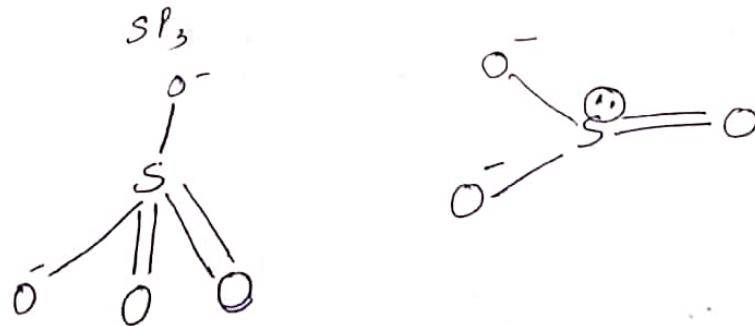
Eg ① $d_{C-O} > d_{N-O}$

Rule ③ → If central atoms & surrounding atoms are same

$B-O \uparrow$; $B-S \uparrow$; $B-L \downarrow$

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

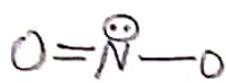
Eg ② $d_{S-O}(SO_4^{2-}) < d_{S-O}(SO_3^{2-})$



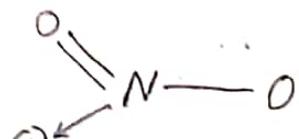
$$BO = \frac{5}{4}$$

$$BO = \frac{4}{3}$$

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



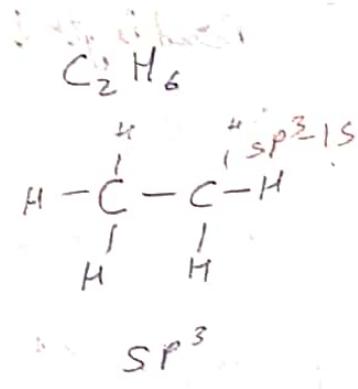
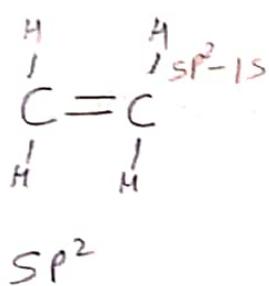
$$B.O = \frac{3}{2}$$



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

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

$$\text{Eg ①. } C_2H_4 <$$

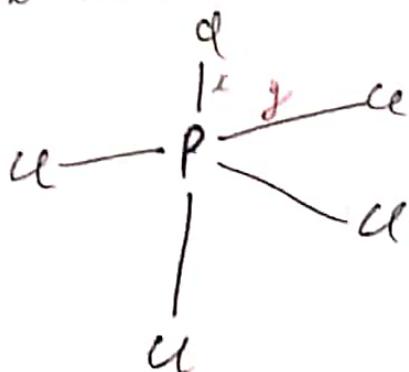


$$\boxed{\% S \uparrow \quad B.O \downarrow}$$

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

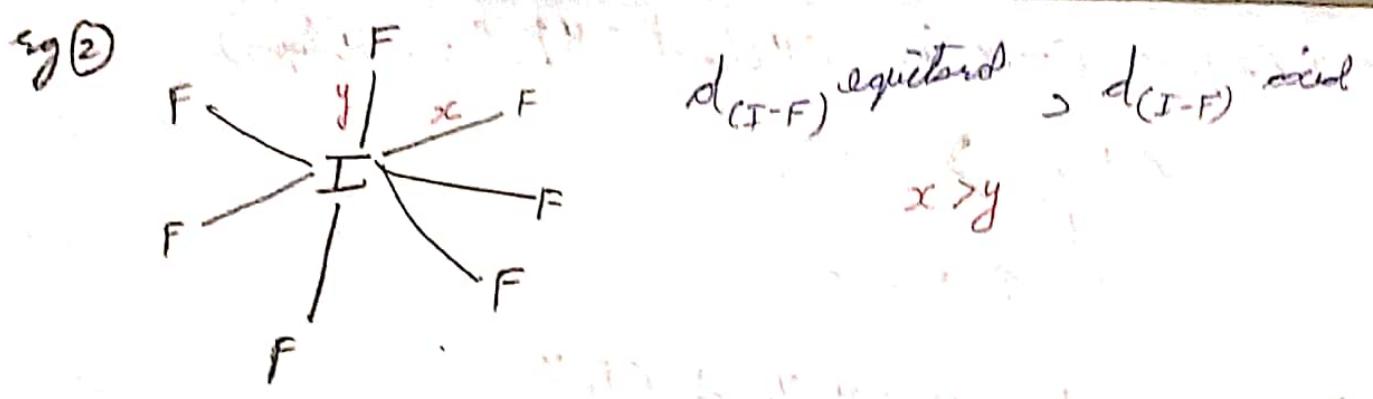
Rule - ⑤ → If central atom, surrounding atom, Bond orders and hybridisation are same then Bond length is decided by Bent's rule.

$$\text{Eg ①}$$

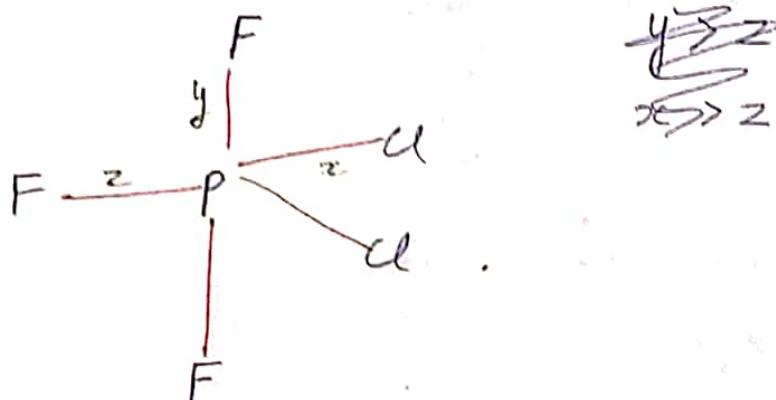


$$d_{(P-Cl)} \text{ axial} > d_{(P-Cl)} \text{ equatorial}$$

$$x > y$$



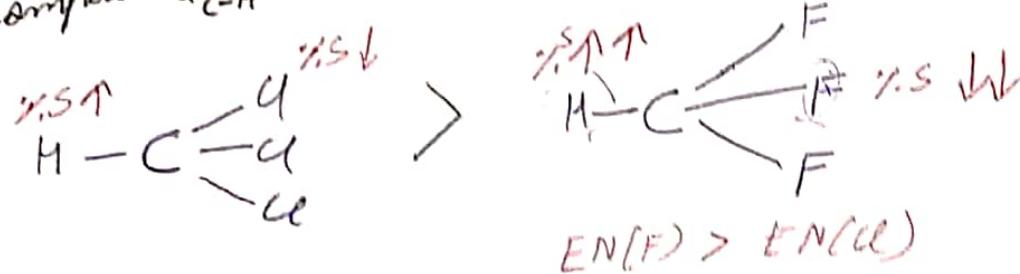
Q PF_3Cl_2



- $d_{P-F}(\text{e}) < d_{P-F}(\circ) \quad z < y$
- $d_{P-F}(\text{e}) < d_{P-Cl}(\text{eq}) \quad z < x$
- $d_{P-F}(\circ) < d_{P-Cl}(\text{eq})$

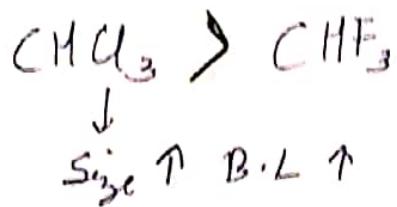
Bent *Caro* *Bhad me*
(size is always dominating)

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



~~CHCl₃~~ $y.s.$ ~~Y.S.T~~ ~~Y.S.T~~, $B.L$ ~~Y.S.T~~ ~~Y.S.T~~

Q2. Compare C-H in CHCl_3 & CHF_3 (\rightarrow Halogen)



Q3. compare $d_{\text{C-H}}$ in CH_3Cl & CF_3Cl

