

Chemical Bonding

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

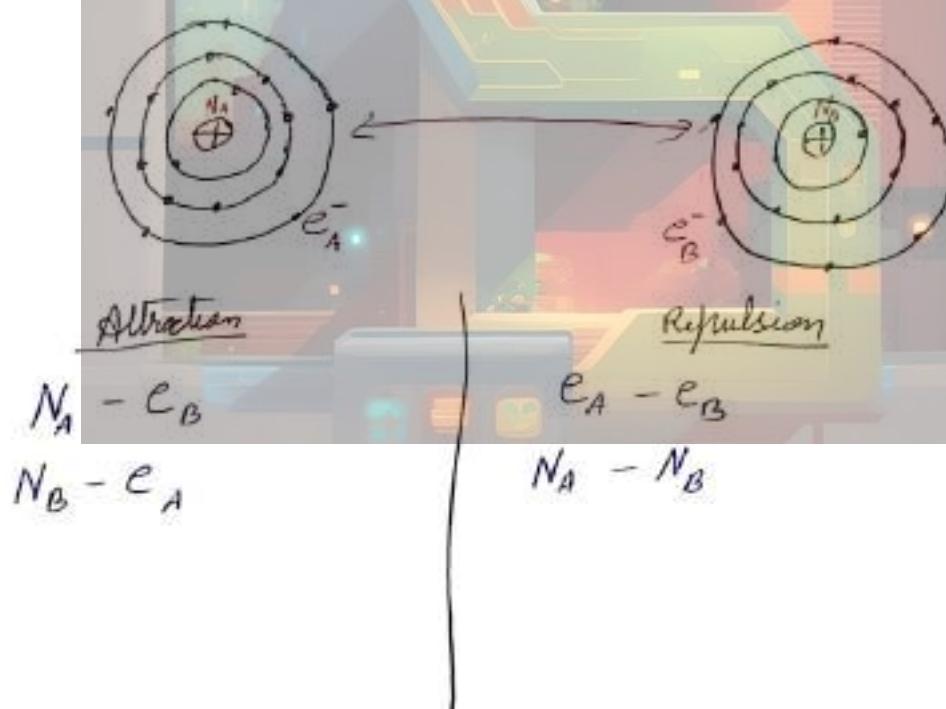
Why a chemical bond is formed?

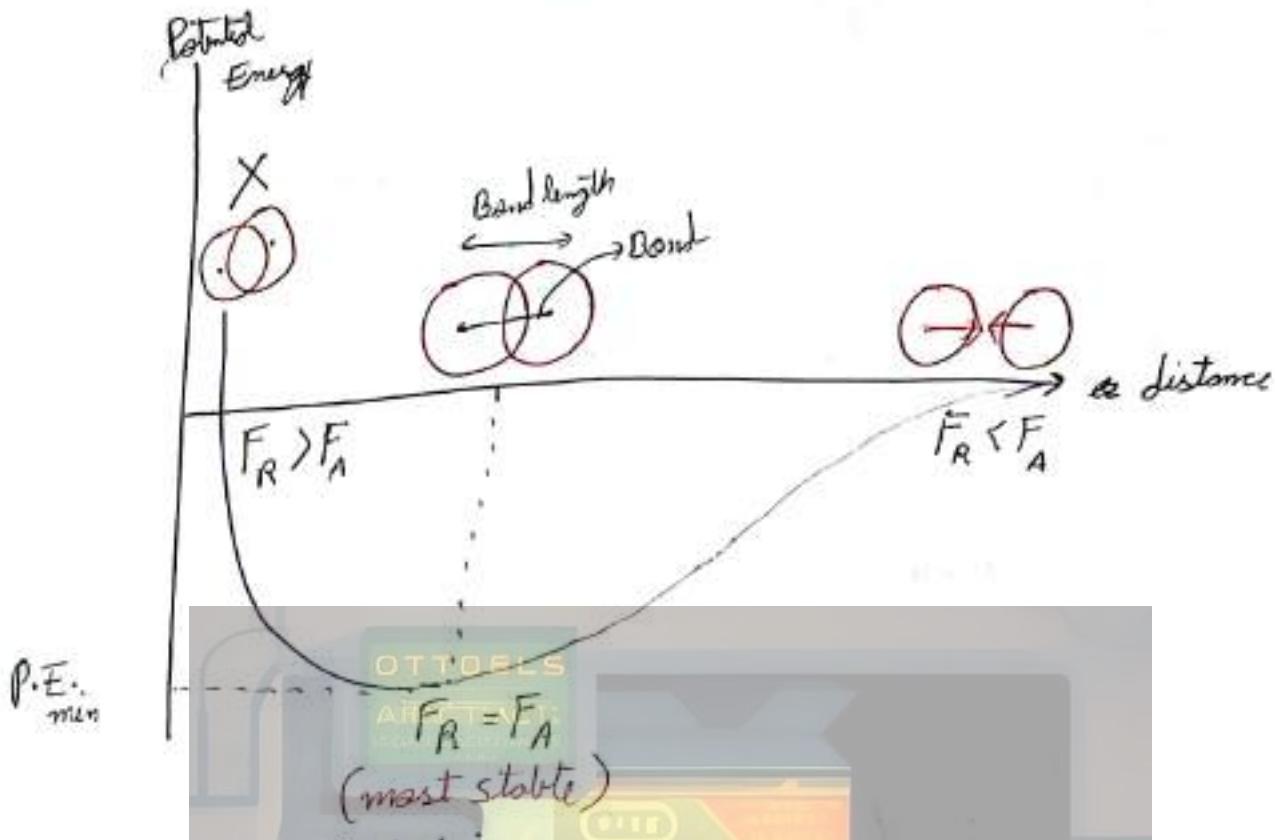
→ To gain stability to do by decreasing their potential energy or to acquire inert gas configuration.

→ 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)





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.

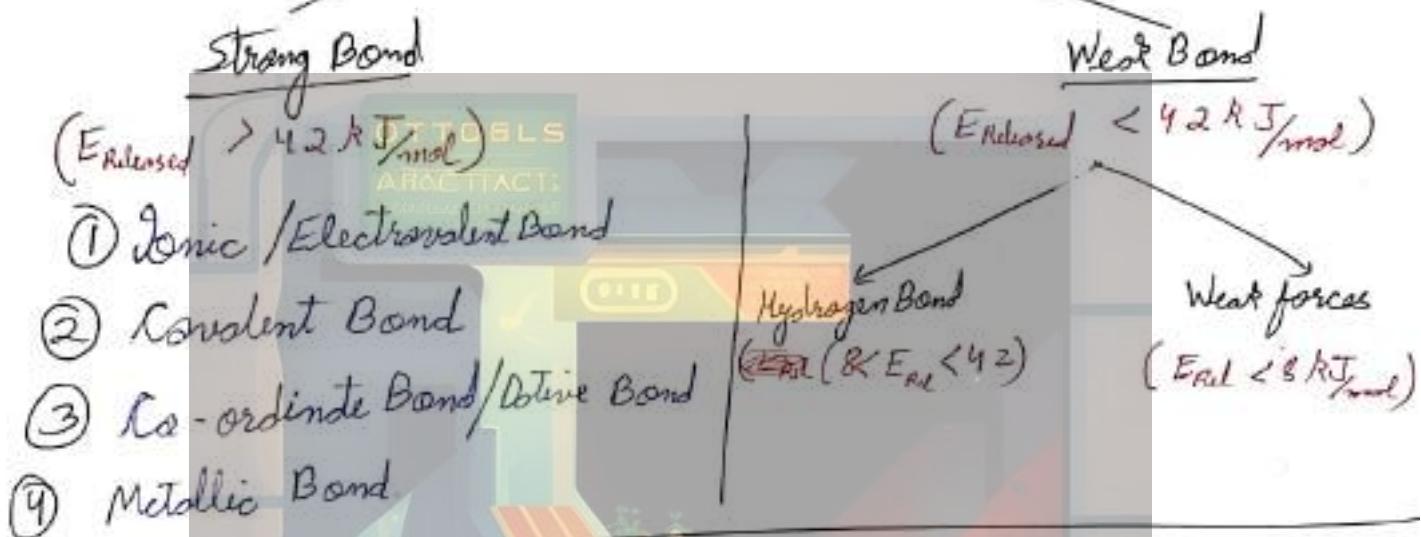
Bond Energy

- It is ~~the~~ the amount of energy required to break bond.
- or
- 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.

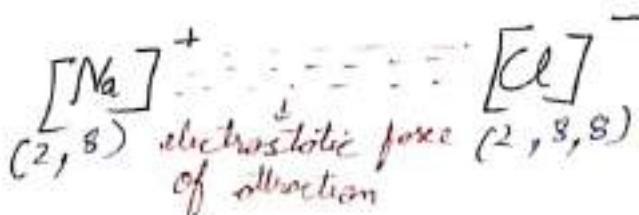
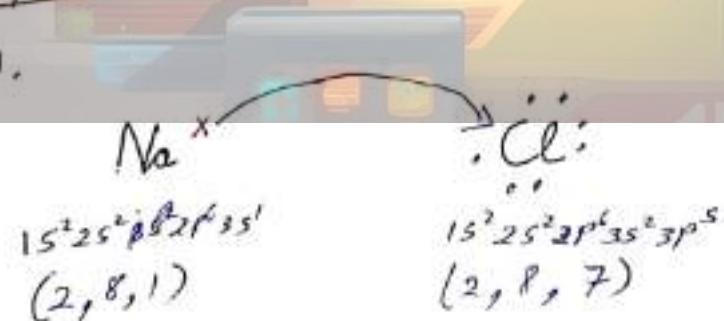
Types of Chemical Bonds.



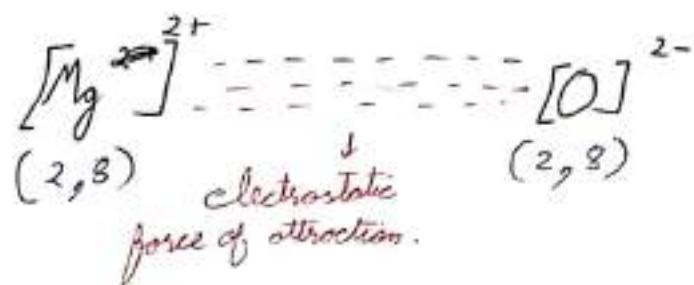
1) Ionic / Electravolent Bond :-

→ 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) = 2

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

3. AlF_3 E.V. (Al) = 3
E.V. (F_3) = 1



Condition for the formation of ionic bond →

$$\rightarrow \boxed{A.E.N \geq 2.1}$$

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

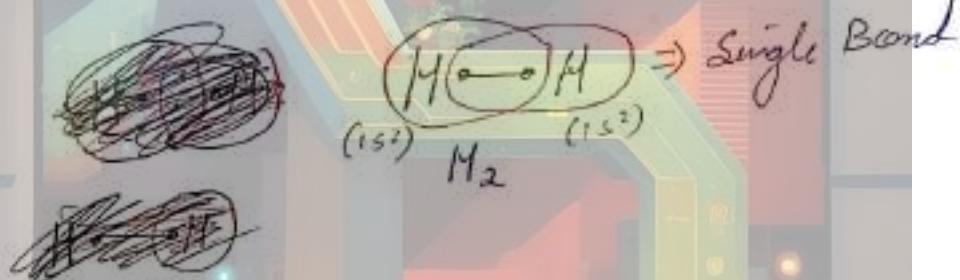
2. Covalent Bond:

→ This type of Bond is formed by equal sharing of electrons between two bonded atoms.

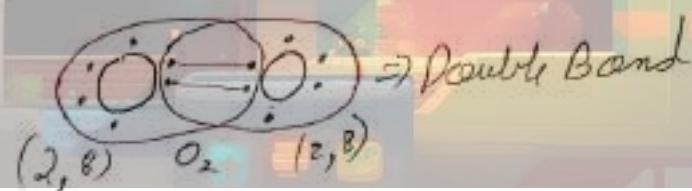
→ ~~Covalent Bonds are~~

→ Covalent Bonds are directional in nature.

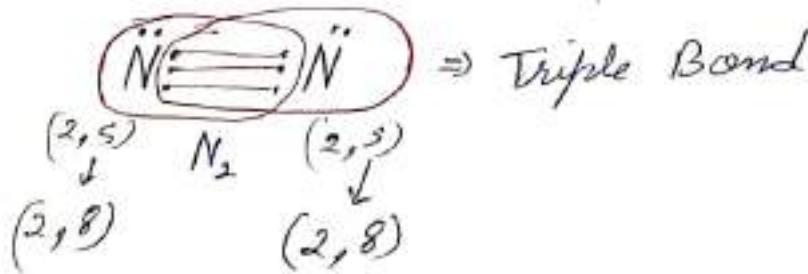
→ Eg 1.



Eg 2.



Eg 3.



Electron Pair

→ Pair of e^- is known as electron pair

Electron Pairs

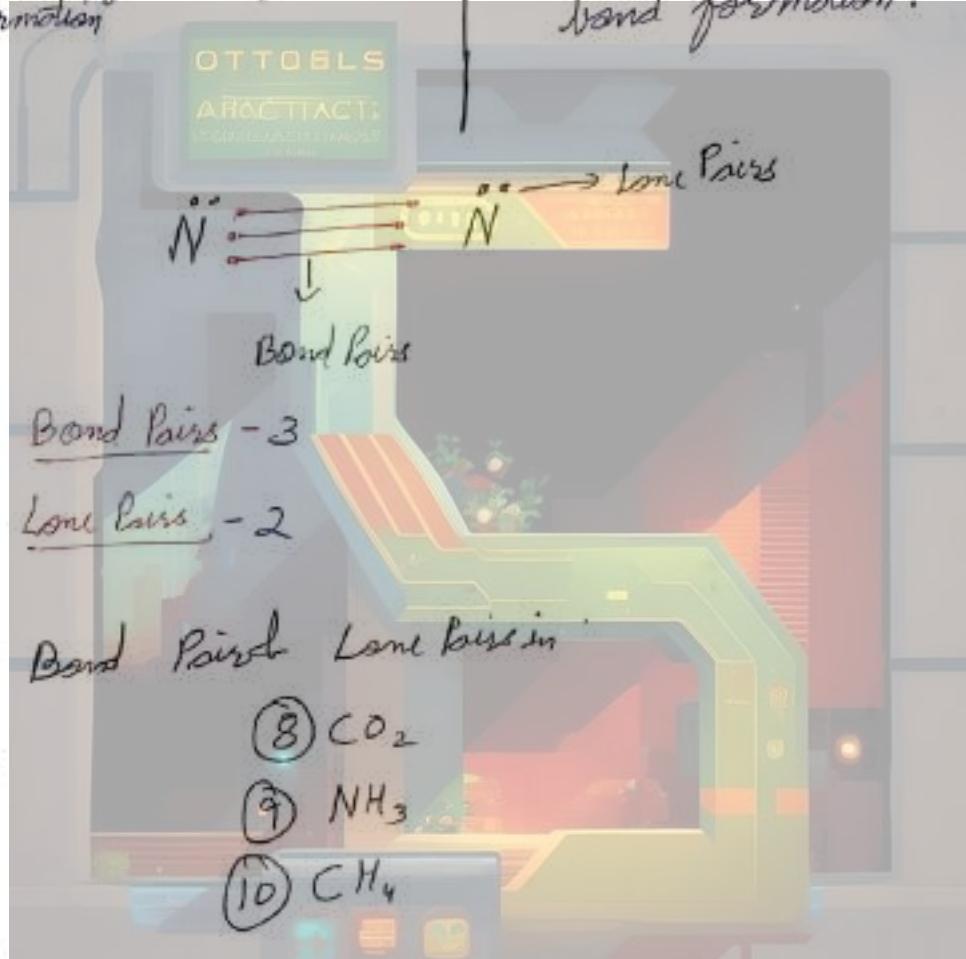
Bond Pairs (B-P)

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

Lone Pairs (L-P)

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

Eg:-



- Q Count
- ① Cl_2
 - ② H_2O
 - ③ CS_2
 - ④ SF_2
 - ⑤ HCl
 - ⑥ NCl_3
 - ⑦ C_2F_2

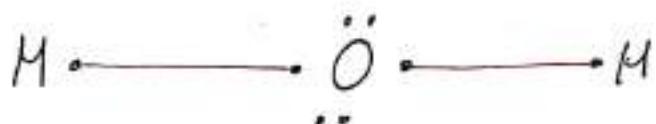
①



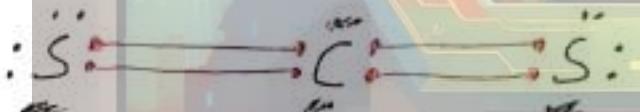
Bond Pair $\rightarrow 1$

Lone Pair $\rightarrow 6$

②

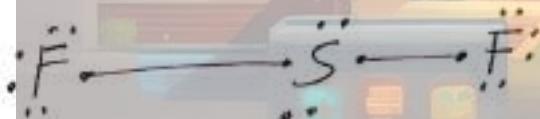


④



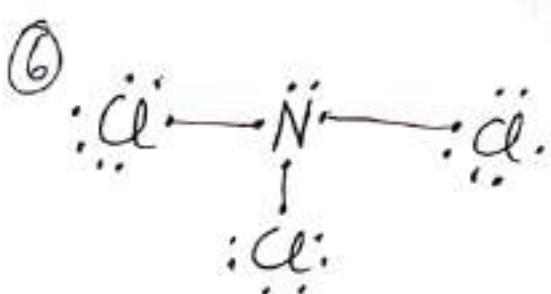
BP - 4
LP - 6 4

⑤

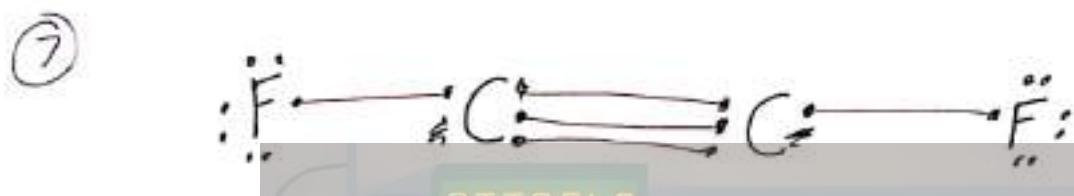


BP - 2
LP - 2 8

(180)



BP - 3
LP - 10



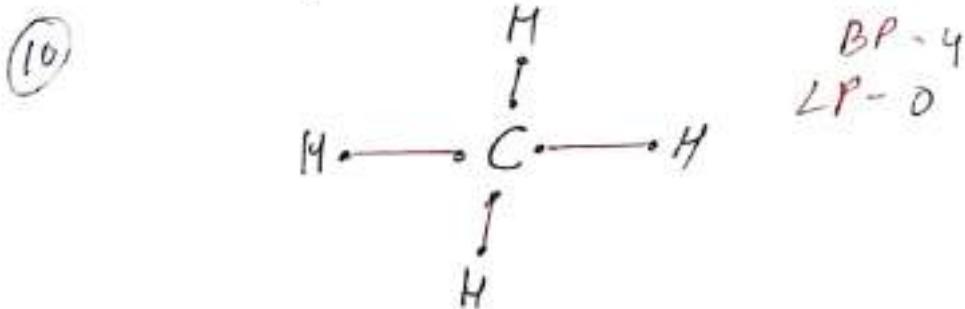
BP - 5
LP - 6



BP - 4
LP - 4



BP - 3
LP - 1



BP - 4
LP - 0

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

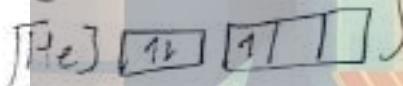
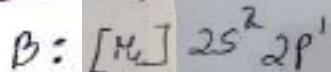
→ It is defined as the number of ^{or} covalent bonds formed by an atom in a covalent molecule ^{or} compound.

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

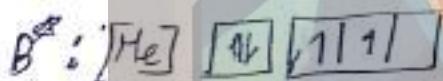
→ In S-block & P-block elements, the number of e^- in outermost shell is called valence electron.

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

Boron Family



$$\} \text{ covalency} = 3 \quad (\text{does not exist})$$



$$C.V. = 3$$

$$\text{covalency} = 3$$

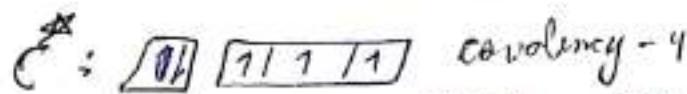
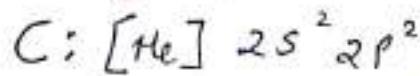
→ The excitation of $1s$ is not possible since shell changes require very high amount of energy.

→ 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)

$$\Delta E N < 2.1$$

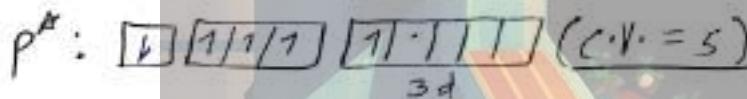
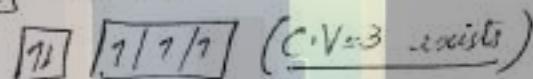
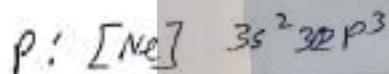
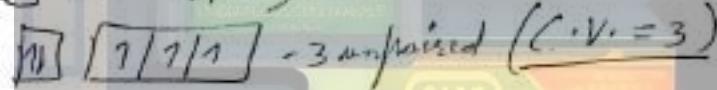
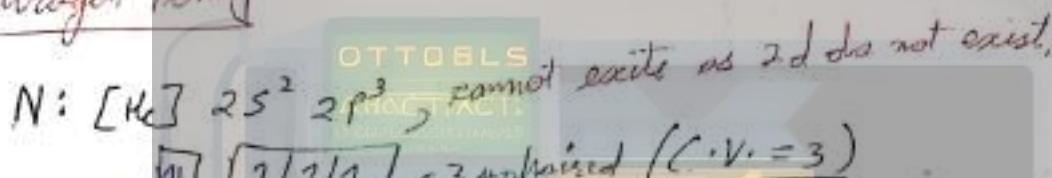
Carbon family - (tetravalent)



Eg①. $CH_4, CO_2, COCl_2, CCl_4$

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

Nitrogen Family



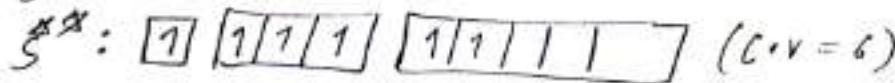
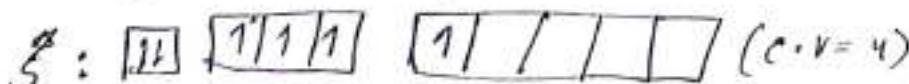
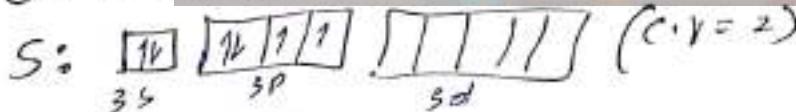
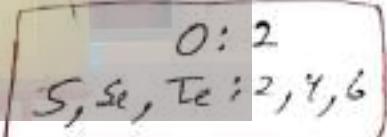
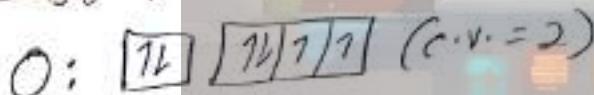
\rightarrow Nitrogen forms only CO_2

\rightarrow Nitrogen has covalency 3.

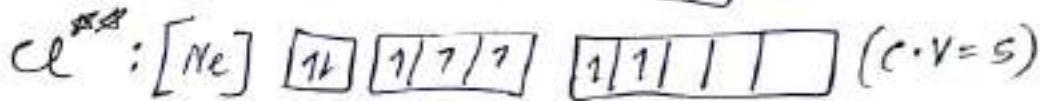
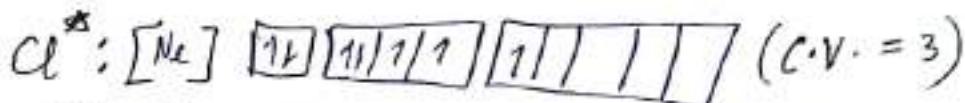
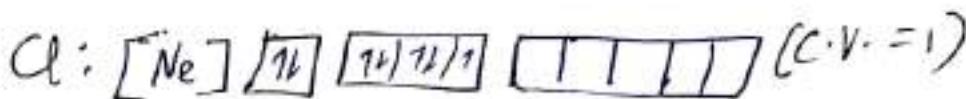
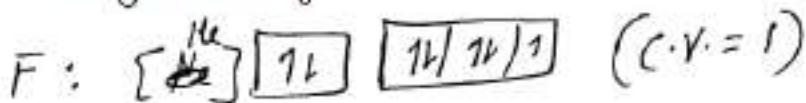
$\rightarrow P, As, Sb$ has covalency both 3 & 5.

Eg① PCl_3, PCl_5

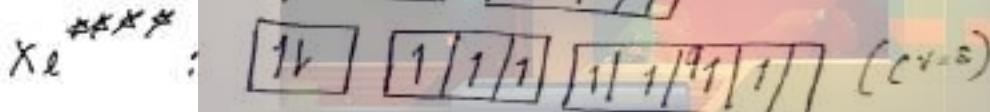
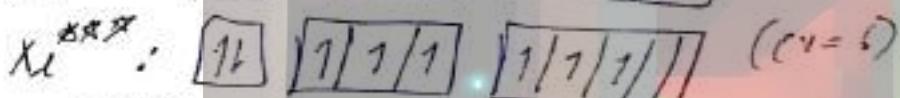
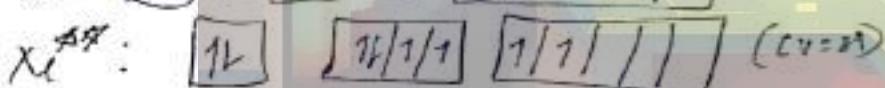
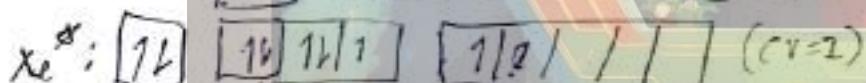
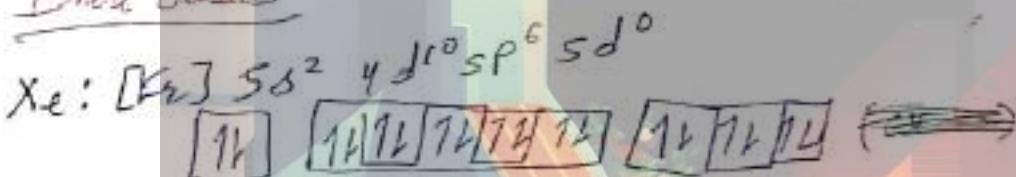
Oxygen family



Molegan Family

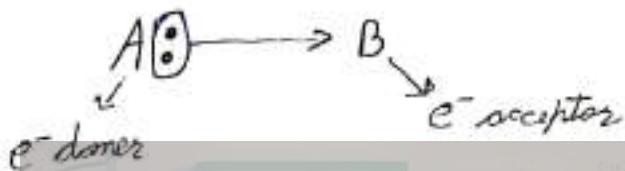


Inert Gases



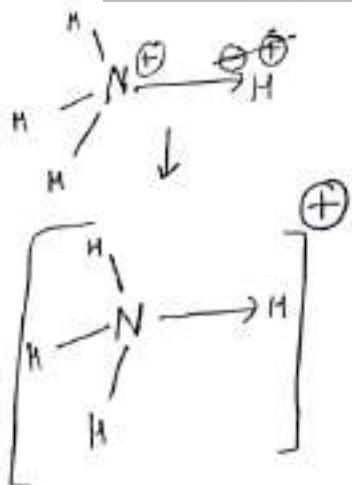
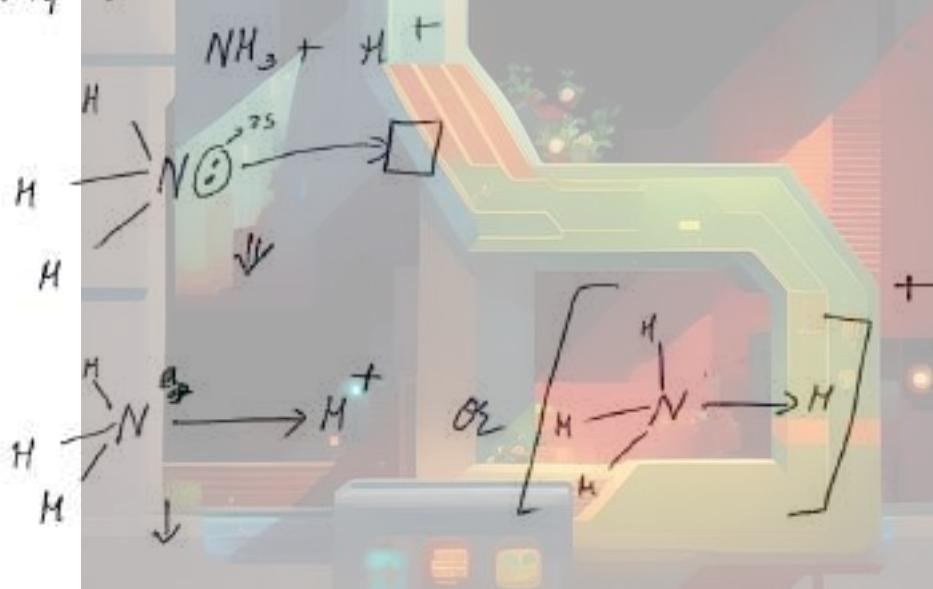
3. Co-ordinate Bond (Dative Bond)

- 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.

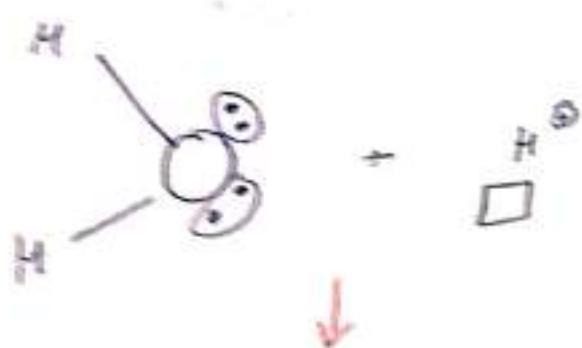


Condition to form co-ordinate bond :-

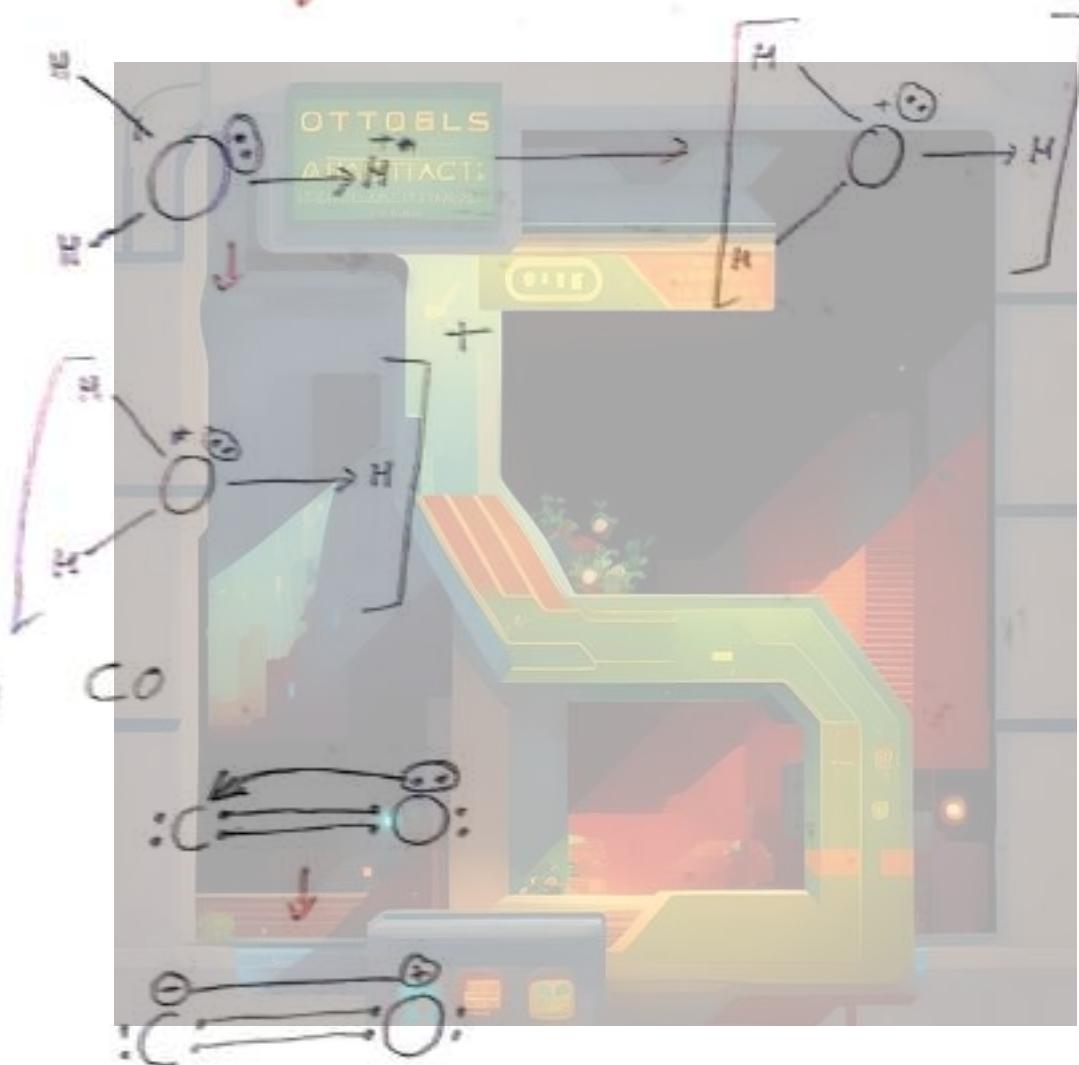
- ① One atom have at least 1 lone pair.
 - ② One atom have atleast one vacant ~~unfilled~~ orbital.
- eg NH_4^+ (Ammonium ion)



Eg ② (H_3O^+)



Eg ③



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



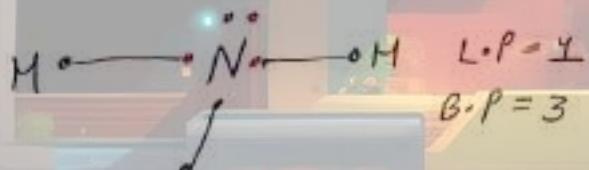
Rules to draw LDS

- ① Close center Central atom -
 → Hydrogen & Fluorine can never be central atom
 → Usually central atom is least electronegative atom and also lesser in number

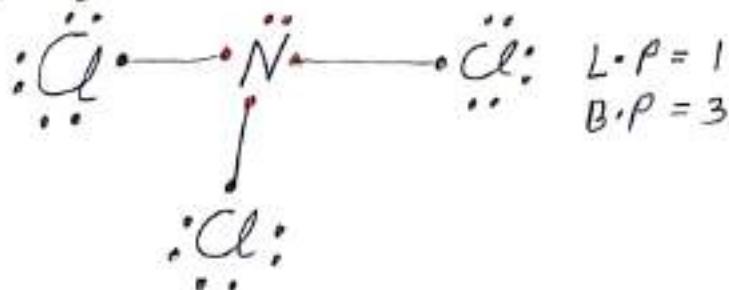
① CO_2



② NH_3



③ NCl_3



④ HCN (Cyanide)

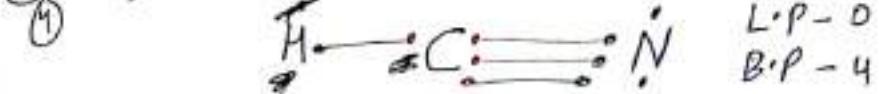
⑤ PH₃

⑥ HNC (Iso-Cyanide)

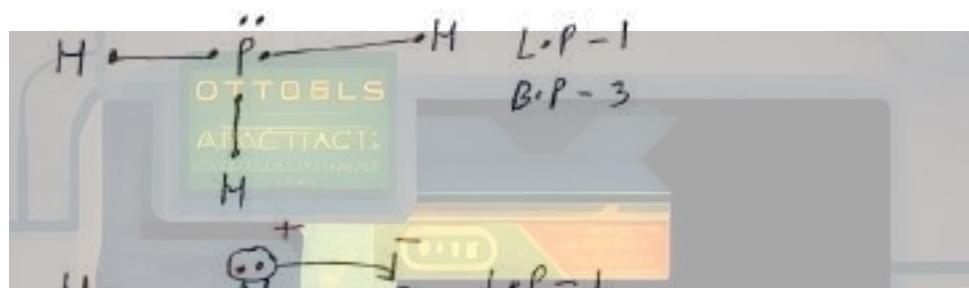
⑦ H₂S

⑧ SO₂

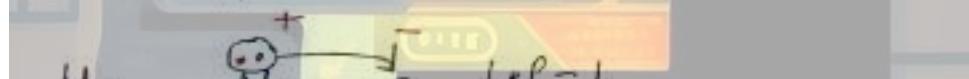
⑨ SO₃



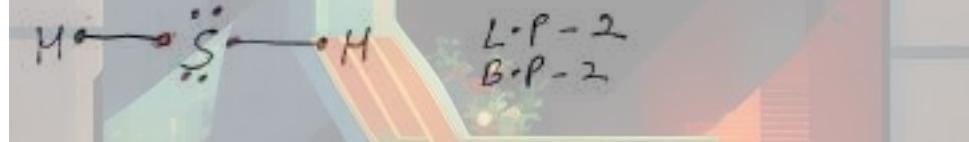
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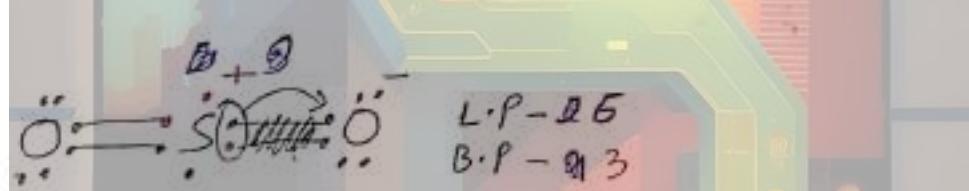
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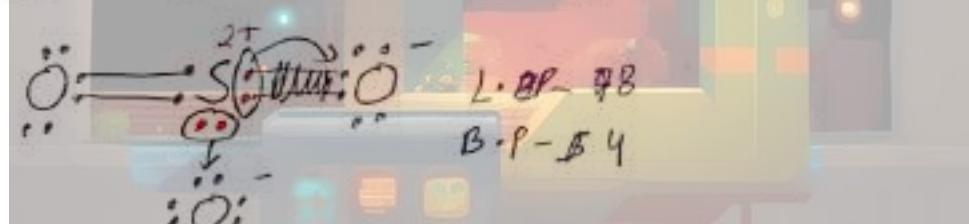
⑦



⑧



⑨



O:

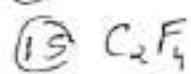
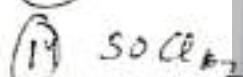
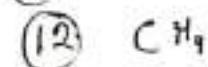
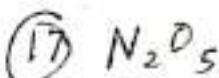
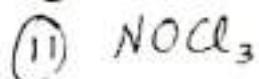
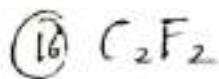
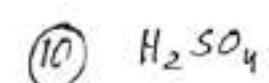
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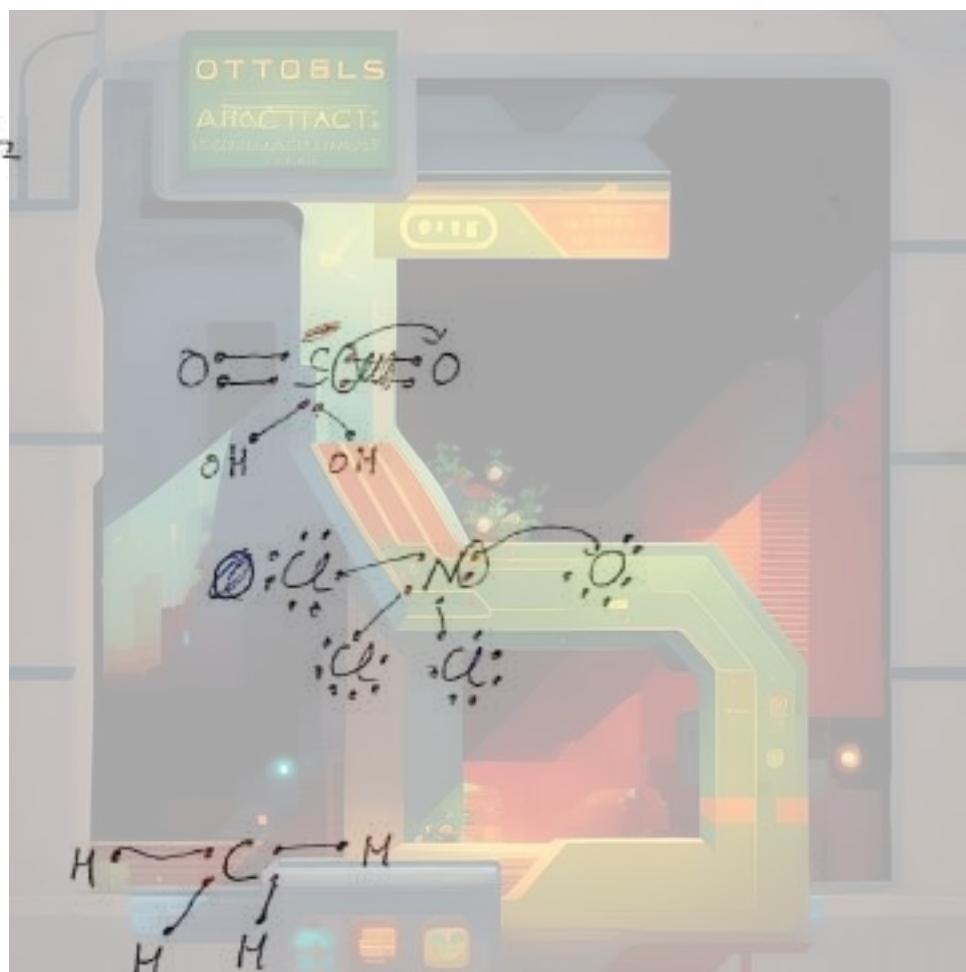
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 → To create vacant orbital for co-ordinate bond.

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.

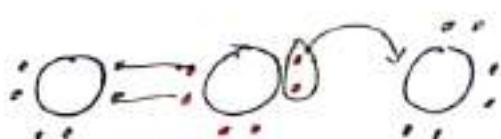


⑩

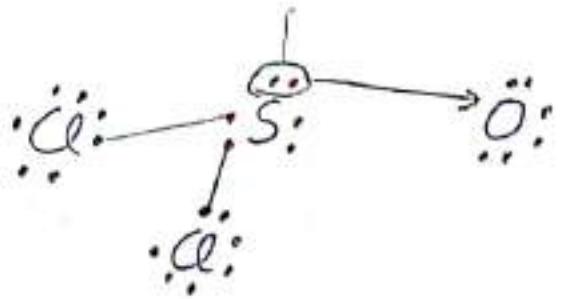


⑪

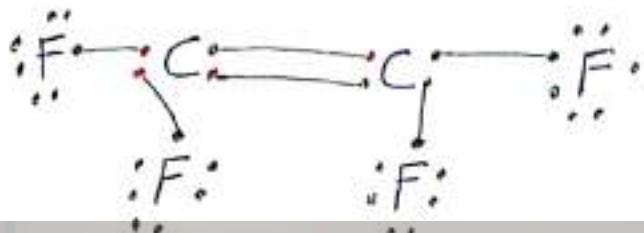
⑫



(14)



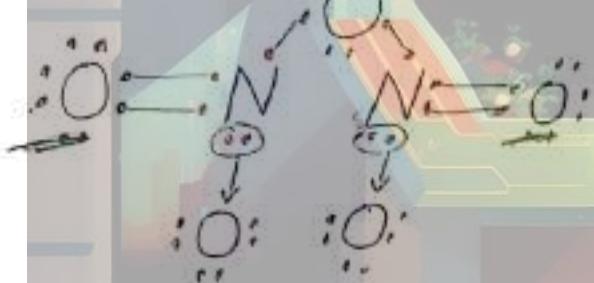
(15)



(16)



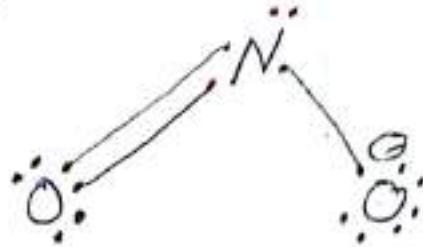
(17)

LDS of charged species

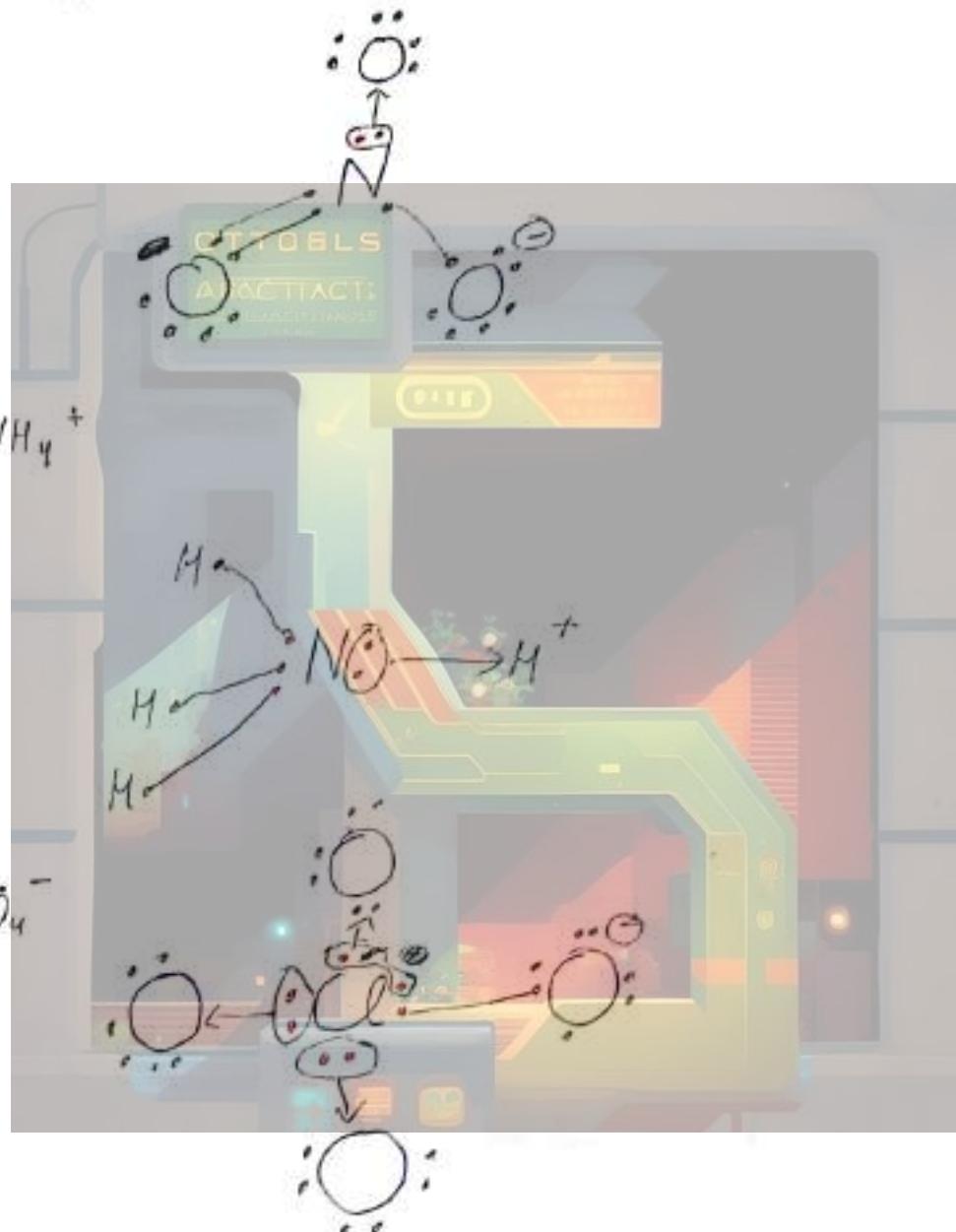
- \ominus ve charge carried by more EN atom
- \oplus ve charge carried by less EN atom.

(18)

Eg ①. NO_2^-



② NO_3^-



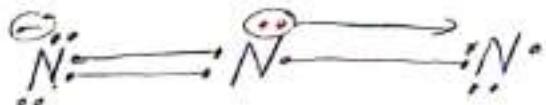
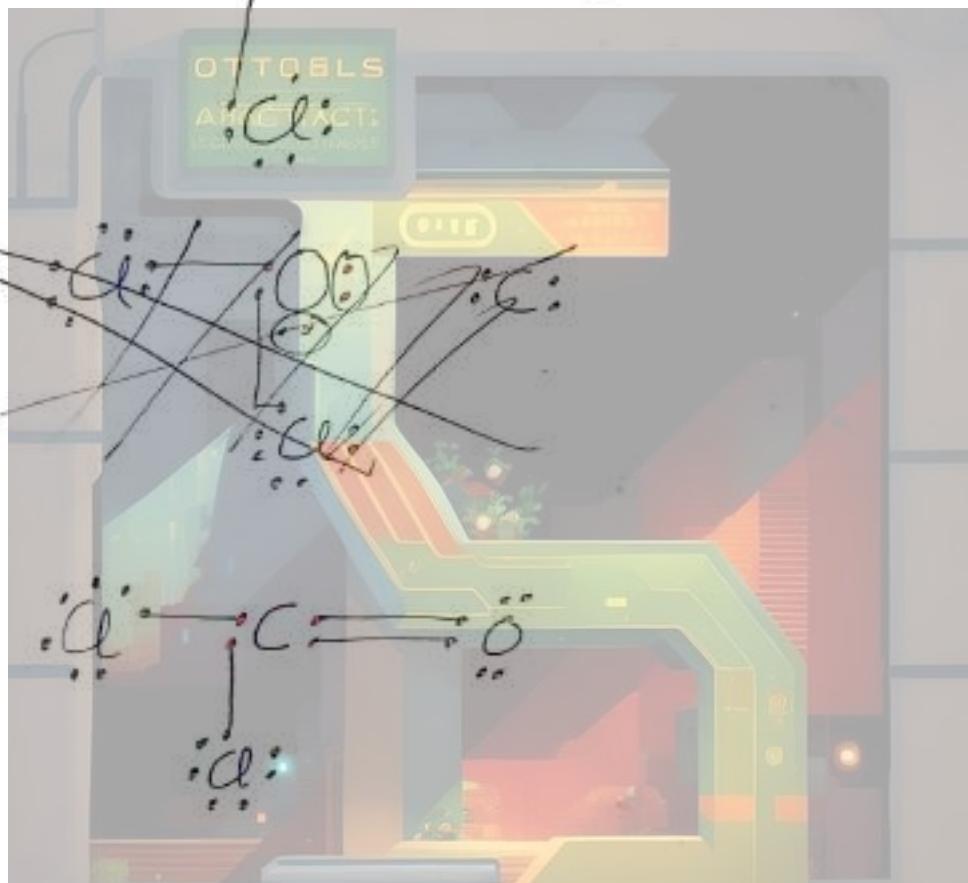
③. NH_4^+

- Q5) Ca^-_3
 Q6) $COCl_2$
 Q7) N^-_3
 Q8) CH_3Cl
 Q9) NH_2Cl

- Q10) OCN^-
 Q11) CN^-
 Q12) SCN^-
 Q13) SiF_4
 Q14) $SnCl_5^-$
- Q15) BF_4^-
 Q16) BH_4^-
 Q17) BeF_4^{2-}
 Q18) CO_3^{2-}
 Q19) NO_3^-Cl

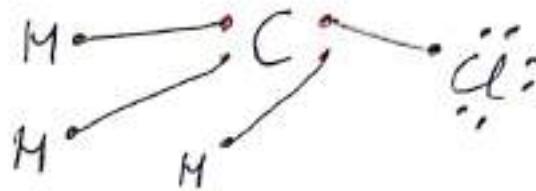
Q20) $NOCl$
 (Tildenreagent)

5)



192

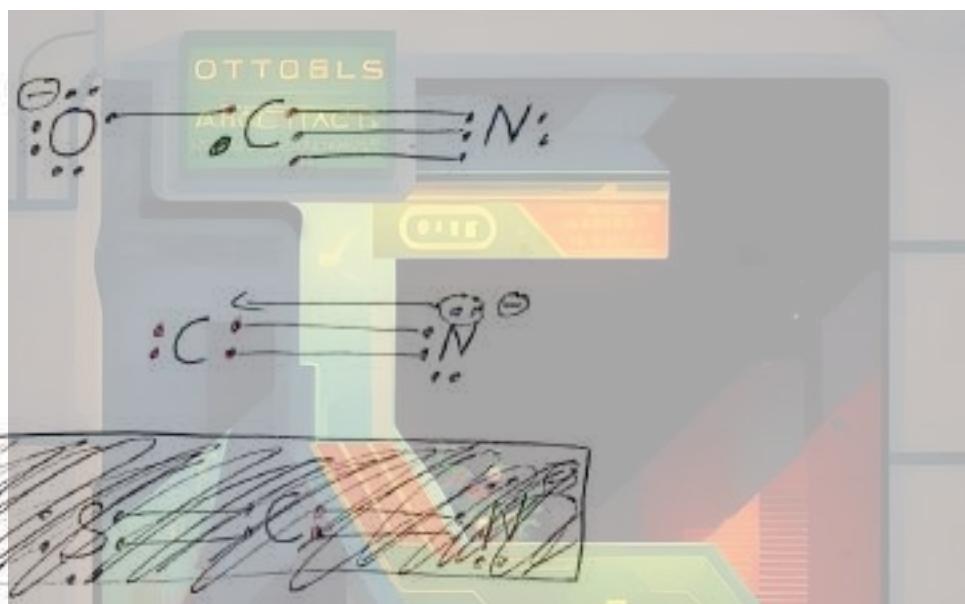
⑧



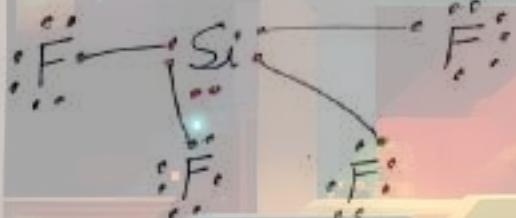
⑨



⑩



⑪

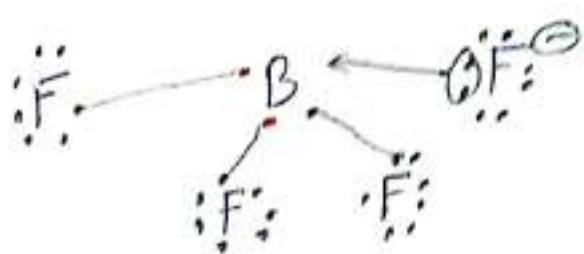


⑫

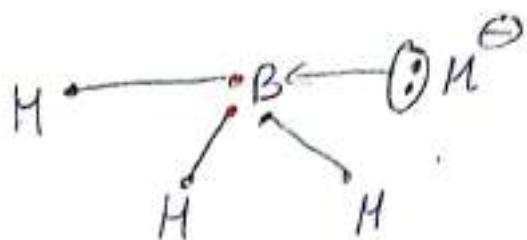


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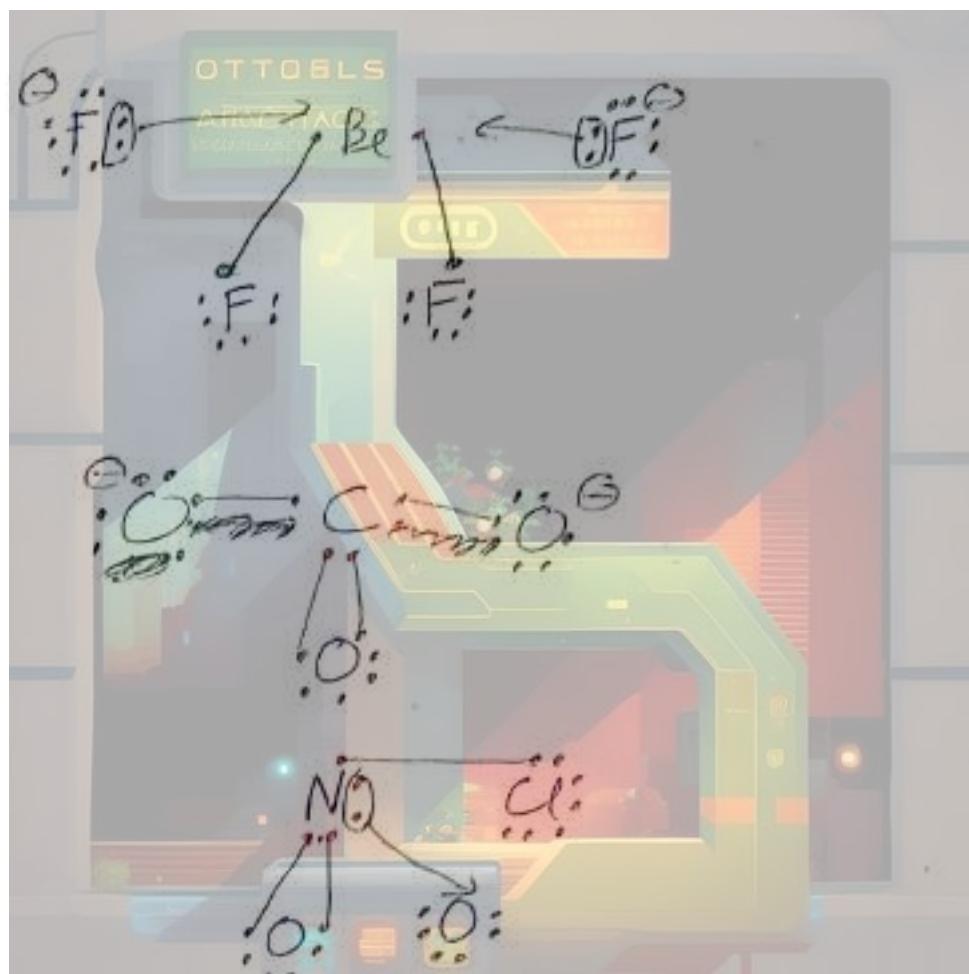
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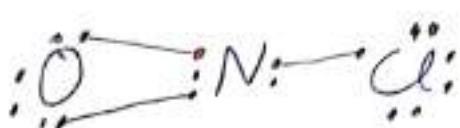


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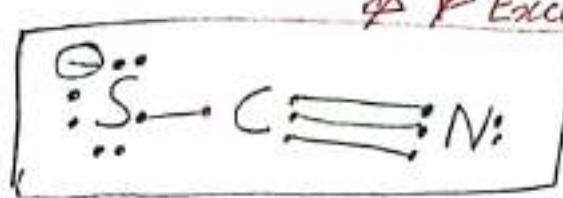
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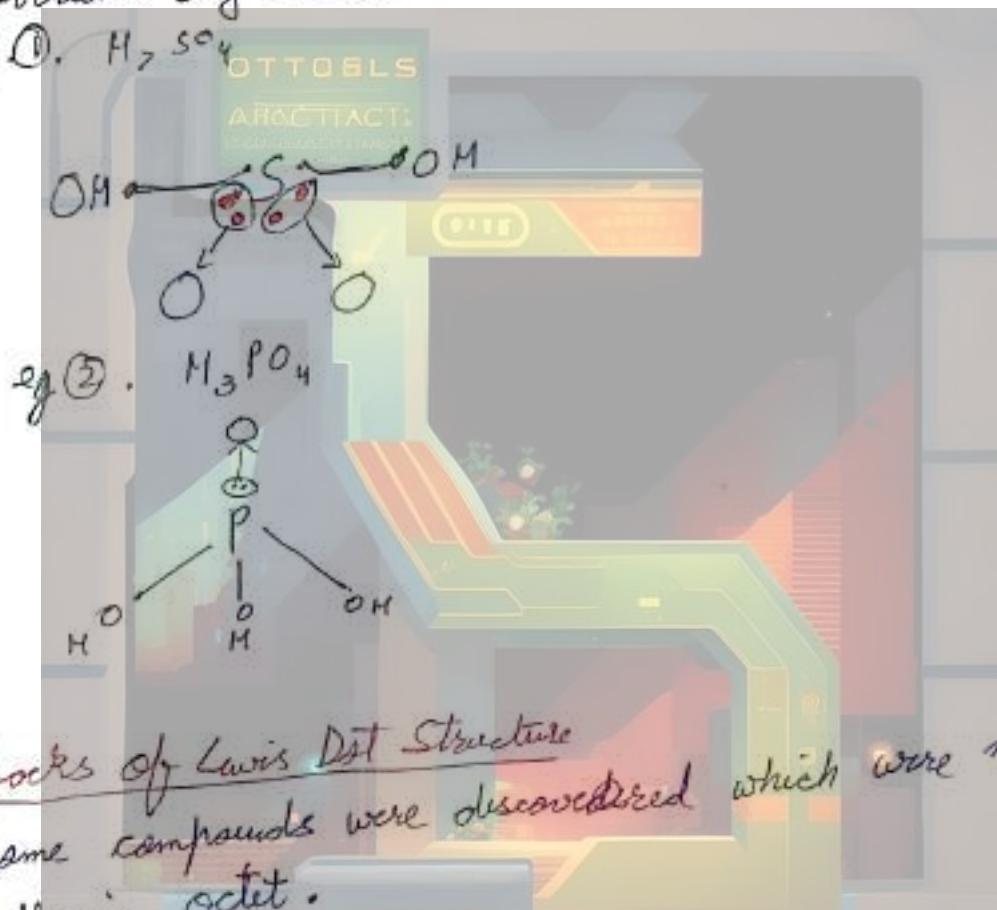
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(12)

Exception

Note:-

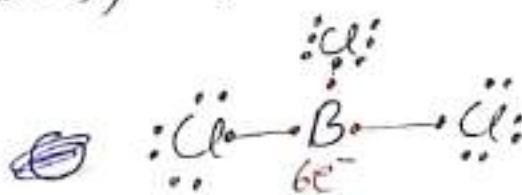
- ① If oxygen and hydrogen both are present together in a molecule. Then connect them with central atom in the form of OH bond. (few exceptions are there will be covered in oxy acids).

eg ①. H_2SO_4 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 , BF_3AlCl_3 

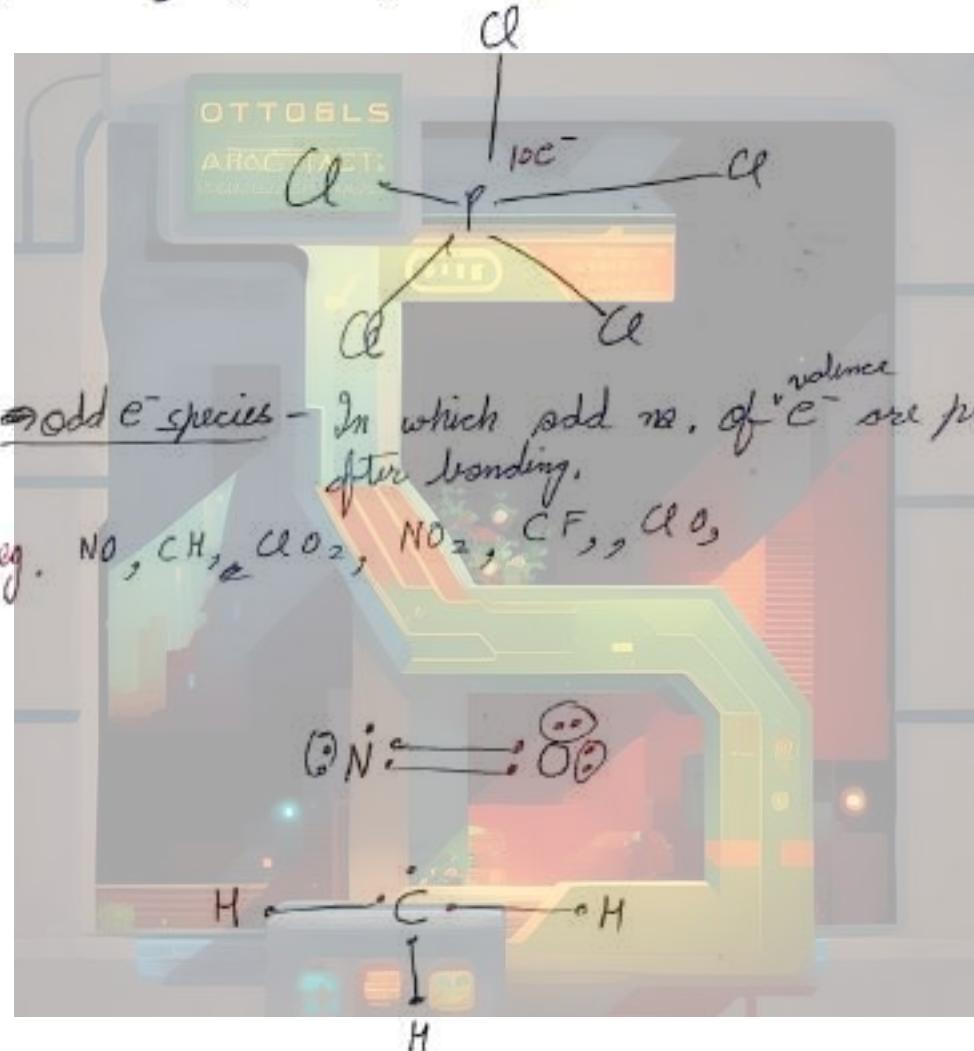
(195)

→ 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. (DEN)

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

e.g. PCl_5 , XeE_2 , 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_3 , ClO_2 , NO_2 , CF_3 , ClO_3



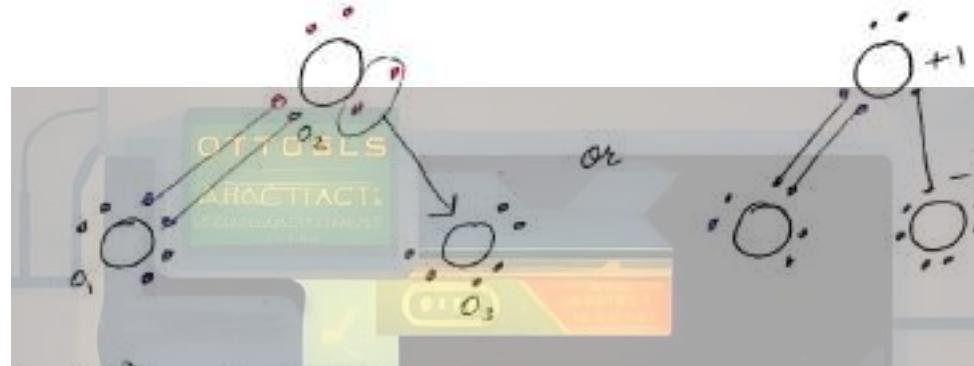
② Lewis theory cannot explain shape of or geometry of molecule.

Formal Charge :-

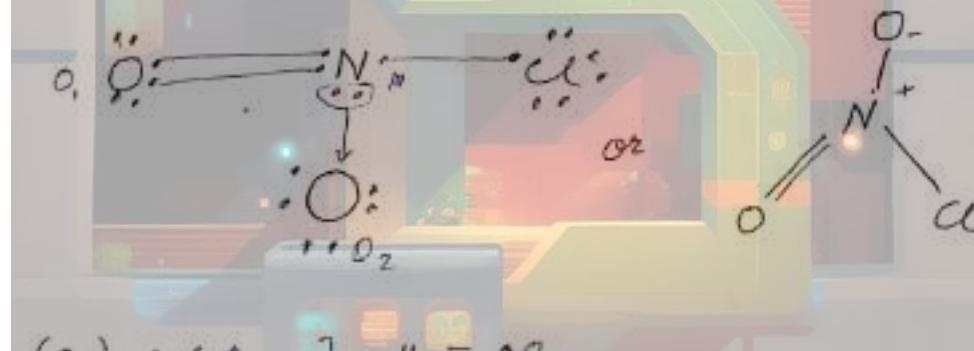
- It is defined for Lewis dot structures.
- It is the charge on atom in a given molecule or ion.

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

Eg ①

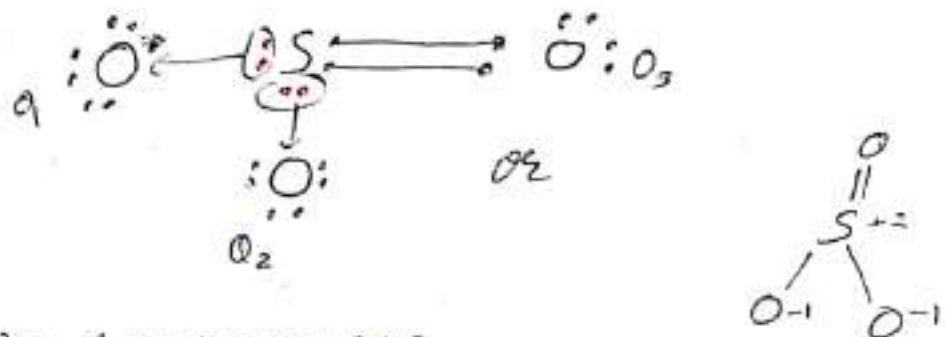


②. $NO_2 Cl$

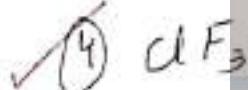
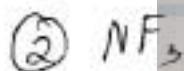
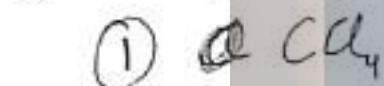


$$\begin{aligned} F.C(O_1) &= 6 + -2 - 4 = 0 \\ F.C.(O_2) &= 6 - 1 - 6 = -1 \\ F.C.(N) &= 5 - 3 + 4 - 0 = +1 \\ F.C.(Cl) &= 7 - 1 - 6 = -1 \end{aligned}$$

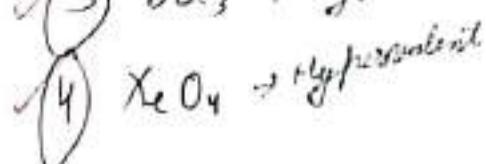
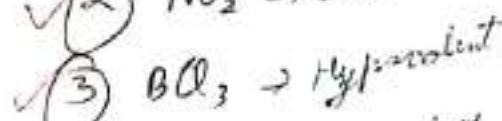
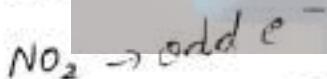
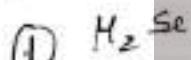
③ SO_3



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



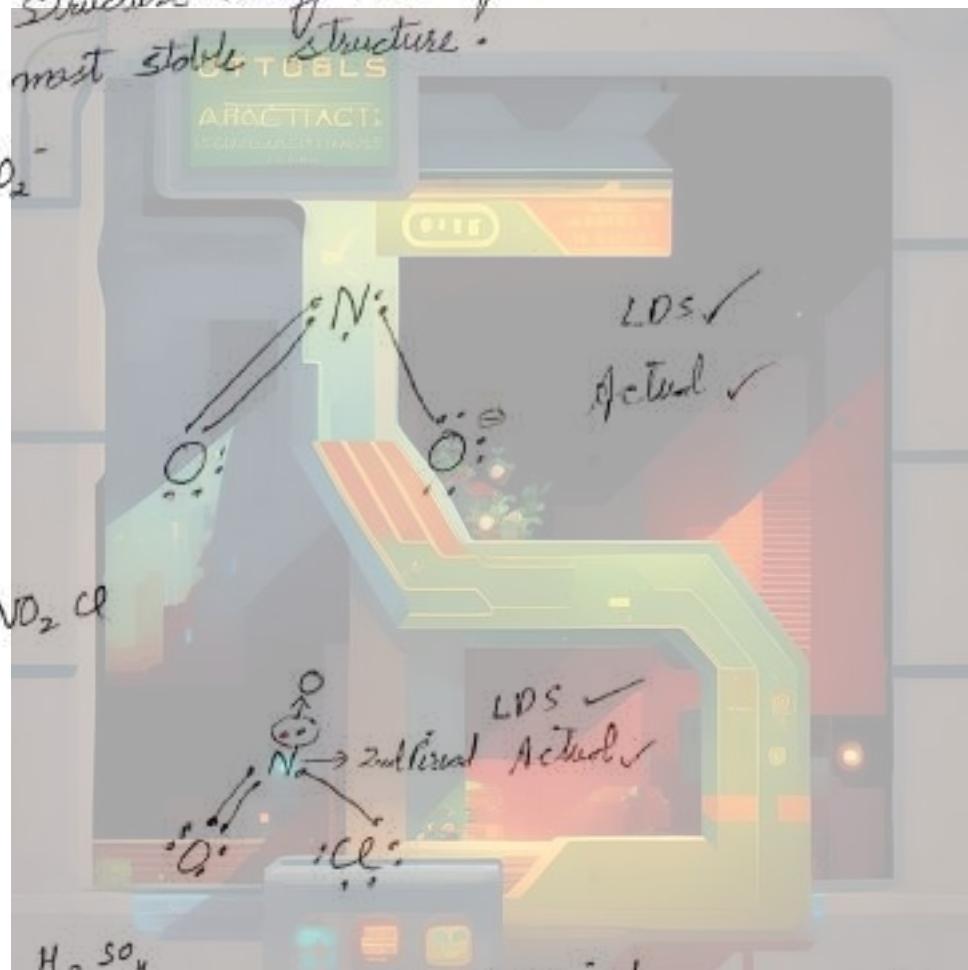
Q. Which of the following species do not follow octet rule?



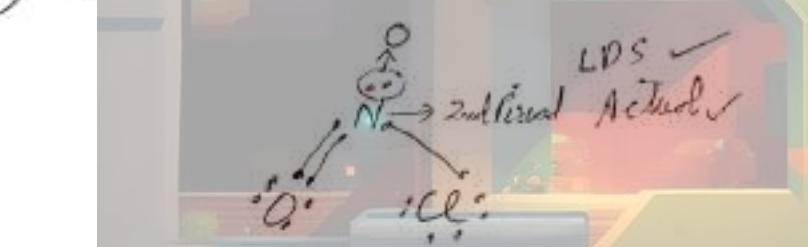
②

- Notes -
- ① All second period elements have Lewis dot structures come 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 ~~for~~ 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.

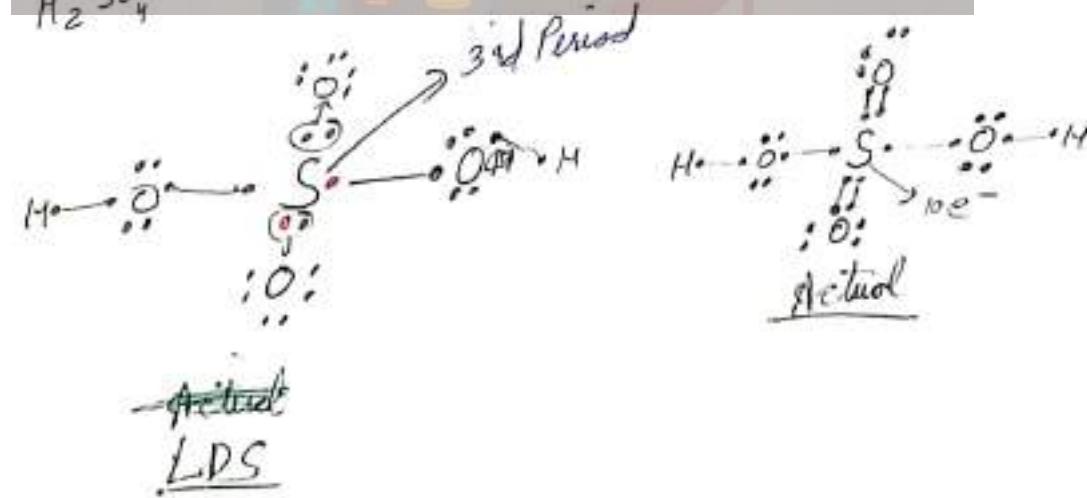
Ex ① NO_2^-



② NO_2Cl

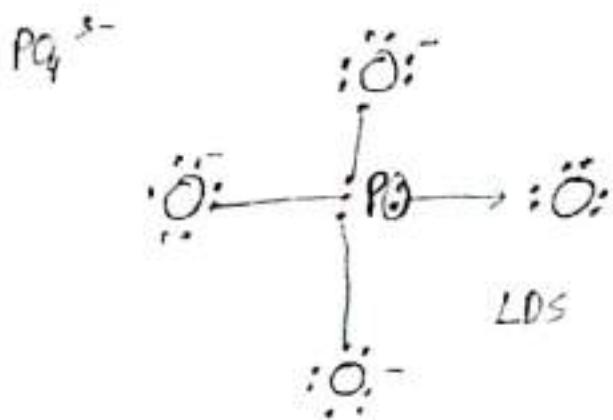


③ H_2SO_4

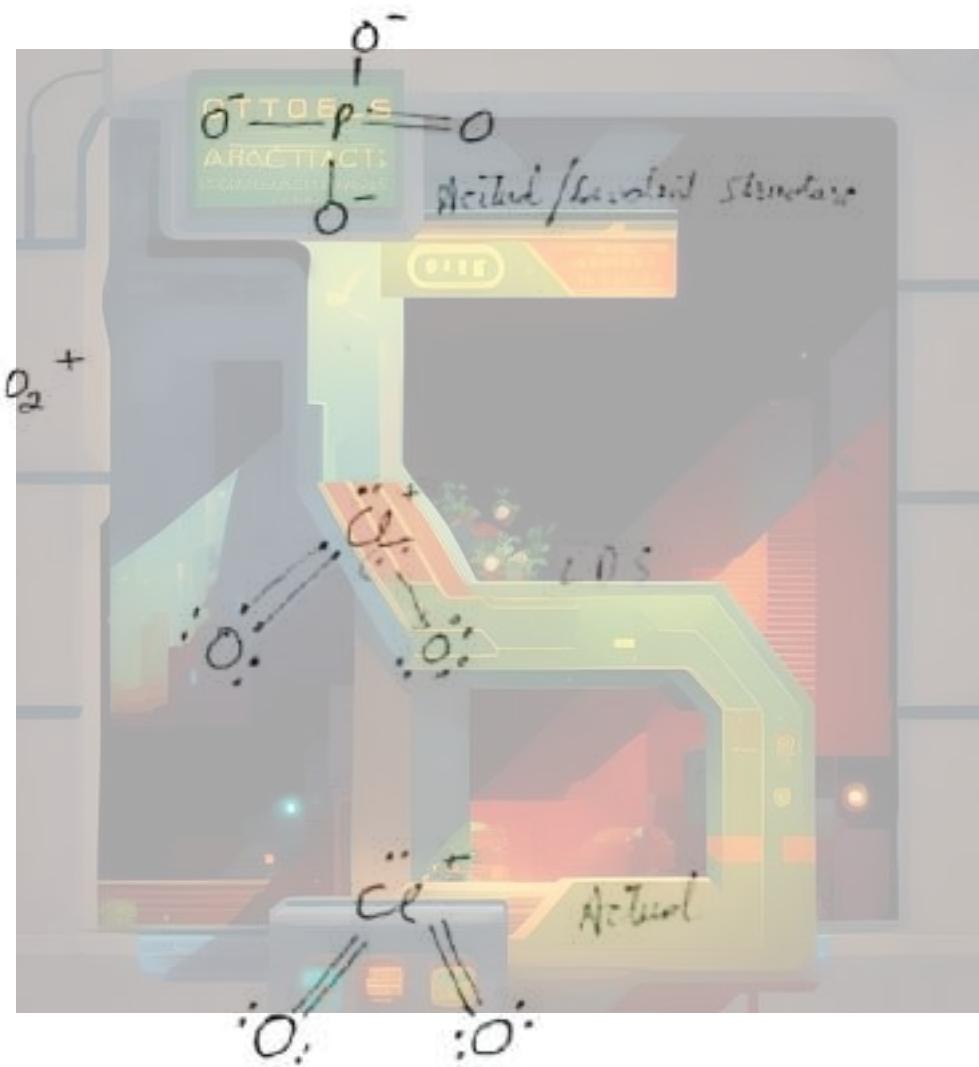


③

④



⑤



⑥

Q Calculate Total no. of L.P.

① SF_6

② SF_4

③ XeF_3

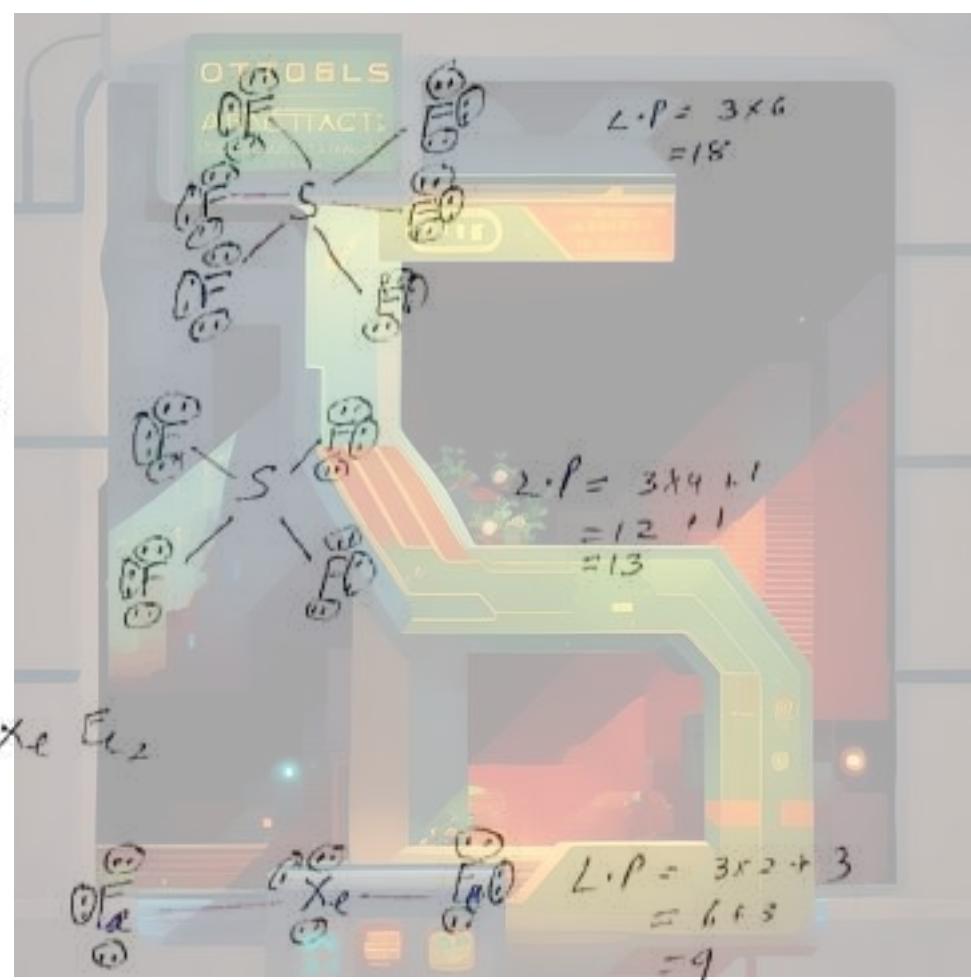
④ XeF_4

⑤ SO_2

⑥ SO_3

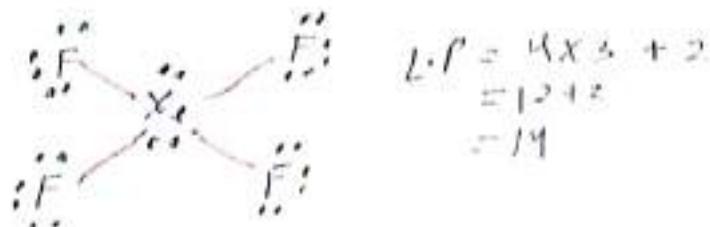
⑦ $NOCl_3$

① SF_6



③ XeF_3

④ XeF_4

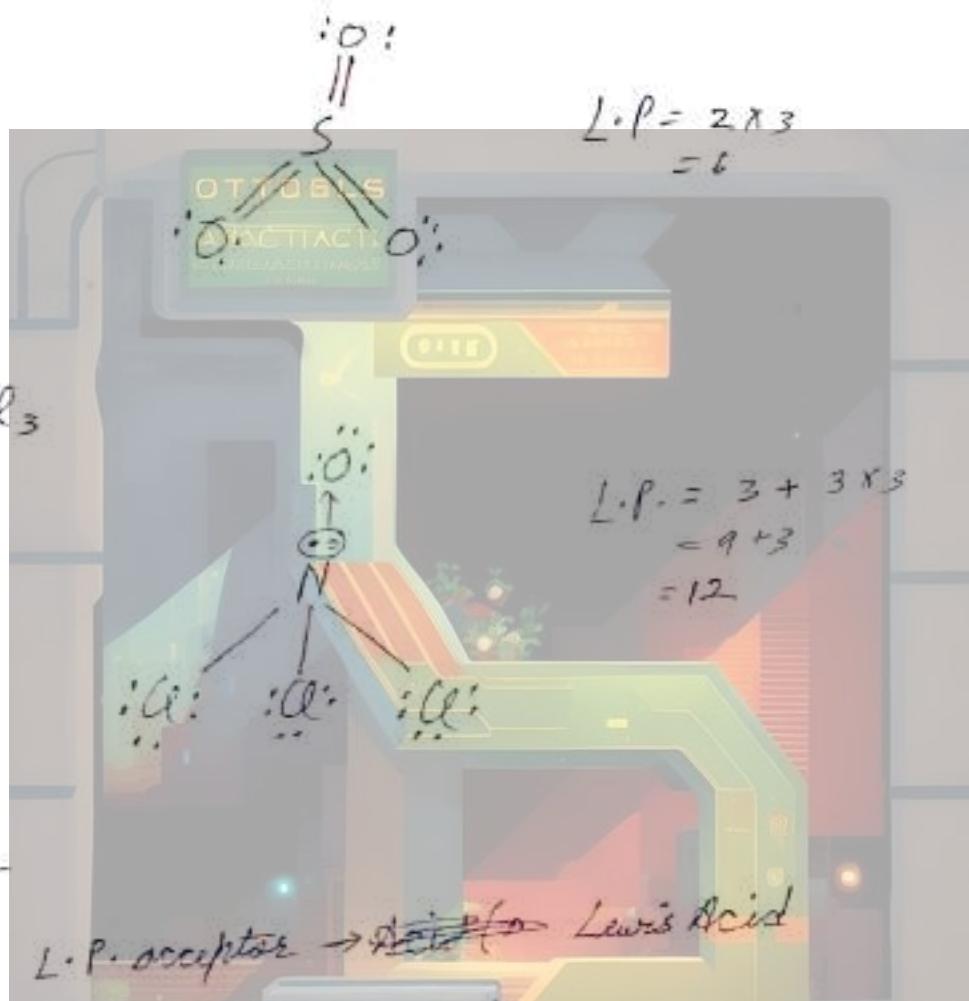


⑤ SO_2



$$\begin{aligned} L.P. &= 6 \times 2 + 2 \\ &= 12 + 2 \\ &= 14 \end{aligned}$$

⑥ SO_3



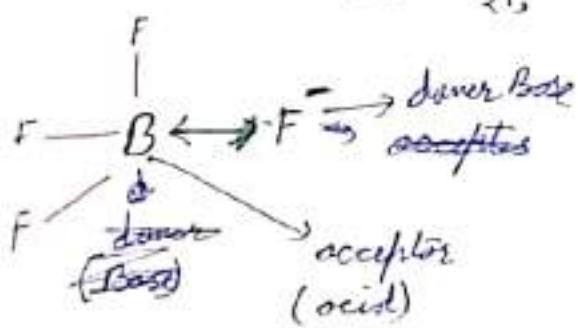
⑦ NO_2

Note :-

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

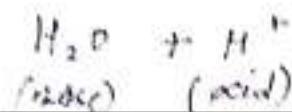
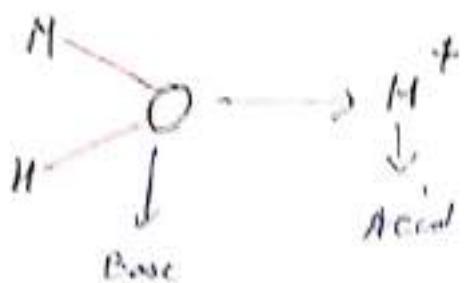
② L.P. Donor \rightarrow ~~Acceptor~~ Lewis Base

③ BF_4^- $B: \frac{1s^2}{1s^2} \quad \frac{2s^2}{2s^2} \quad \frac{2p^3}{2p^3} \rightarrow$ vacant orbital



⑥

② H_2O



OTTOELS
ABSTRACTS
1950-1951

Oxidation Number -

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

→ Rules for calculating oxidation number.

① Oxidation no. of free molecule or pure element is zero.

e.g. Na , Cl_2 , N_2 , O_2 , H_2 , Ar , Zn

→ Oxidation no. of alkali metals in compound state.

② Oxidation no. of alkali metals (Li , N , K , Rb , Cs , F) $\Rightarrow +1$

e.g. $LiCl$, $NaCl$, Na_2SO_4

→ Oxidation no. of alkaline earth metals in compound state.

③ Oxidation no. of Alkaline earth metals (Ba , Mg , Ca , Sr , Ba) $\Rightarrow +2$

e.g. $Ca(OH)_2$, $Mg(O)$, $Be(O_2)$, $Sr(OH)_2$, CaO

④ Oxidation no. of H :

→ If bonded with metals $\rightarrow -1$

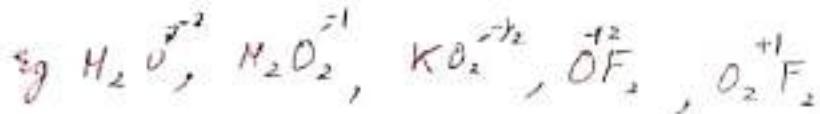
→ If bonded with other element/non-metals $\rightarrow +1$

e.g. CH_4 , LiH , HNO_3 , Ca^{2+} , H^+ , Cl^-

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 Chlorine in compound state $\Rightarrow -1$.

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

H.W.

Q ① SO_2

$$S + 2(-2)$$

$$S = -4$$

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

② SO_3

③ $KMnO_4$

④ HNO_3

⑤ H_2O_2

⑥ H_2SO_4

⑦ KO_2

⑧ Ca_3N_2

⑨ KNO_2

⑩ KNO_3

⑪ $Al_2(SO_4)_3$

⑫

⑧ SO_4^{2-}

⑨ SO_3^{2-}

⑩ NH_4^+

⑪ Cl^-

⑫ MnO_4^{2-}

⑬ $C_2O_4^{2-}$

⑭ OF_2

⑮ VO

⑯ V_2O_5

⑰ NH_4NO_3

⑱ ZnO

⑲ Fe_2O_3

⑳ Pb_3O_4

㉑ Fe_3O_4

㉒ $Cr_2(SO_4)_3$

㉓ $HClO_4$

㉔ N_2O_3

㉕ NO

㉖ PH_3^+

㉗ $HClO_3$

㉘ $HClO_2$

㉙ $HClO$

㉚ LiO_4^{2-}

㉛ NO_2^-

㉜ AlN

㉝ BeF_4^{4-2-}

㉞ $MgCO_3$

㉟ $NaHSO_4$

㉟ H_2S

㉟ N^-

㉟ Al

㉟ $Na_2S_2O_3$

㉟ PO_4^{3-}

㉟ $H_2PO_4^-$

㉟ HIO_4

㉟ H_3PO_4

㉟ H_3PO_2

㉟ CaH_2

㉟ ClO_3^-

㉟ BF_3

② SD₃

$$s + 3(-2)$$

$$s + -6 = 0$$

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

③ $M_3 = 8 - 1$

$$\boxed{M_3 = 7} \checkmark$$

④ $2s - 8 + 5 = 0$

$$s = 5 - 2$$

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

⑤ ⑦ $s + 1 - 6 + N = 0$

$$N = 6 - 1$$

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

⑥ ⑧ $2 \cdot 0 + 2 = 0 = 0$

$$20 = 2$$

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

⑥ ⑦ $1 + 20 = 0$

$$20 = -1$$

$$\boxed{0=-1/2} \checkmark$$

⑧ $s - 8 = -2$

$$s = 8 - 2$$

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

⑨ $s - 2 = -6$

$$\boxed{s=5}$$

$$s = 6 - 2$$

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

⑩ $N + 4 = 1$

$$N = 1 - 4$$

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

⑪ $\boxed{N+4} \checkmark$

⑫ $M_3 - 8 = -2$

$$\boxed{M_3 = 6} \checkmark$$

⑬ $2G_2 + -14 = -2$

$$2G_2 = 12$$

$$\boxed{G_2 = 6} \checkmark$$

⑭ $D + 2 = 0$

$$\boxed{D = +2} \checkmark$$

⑮ $V - 2 = 0$

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

⑯ $2V - 10 = 0$

$$V = 10/2$$

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

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 $1 + Cl + G = 0$

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

㉒ $1 + Cl - 4 = 0$

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

㉓ $Cl = 1$

$$\boxed{Cl=1} \checkmark$$

Note -

odd/even State

① For any closed range of $\sigma, S \Rightarrow N \text{ to } N+8$ (both included)

$N \Rightarrow$ no of valence e⁻

Eg ① N: $N=5$

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

② F; S: $N=c$

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

③ O: $N=O \text{ TOSLS}$

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

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

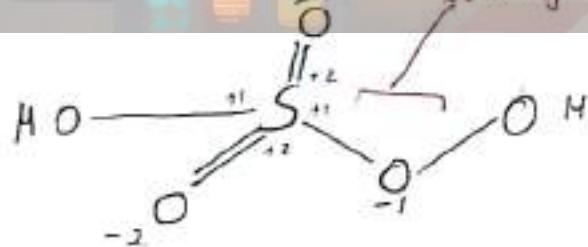
④ Cl: $N=?$

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

Eg H₂S (closed shell solid)

$$2+x-10=0$$

$x = 10 - 2$
 $x = 8$ \rightarrow I had only $\frac{1}{2} \times (-2 \text{ to } 0)$
by writing

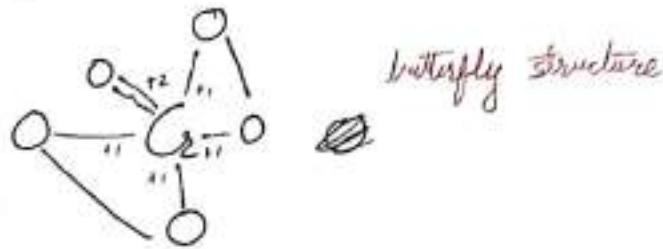


$$\begin{array}{|c} 2+2+1+1=6 \\ \hline O.Sq S=6 \end{array}$$

Eg. CrO_5

$$x - 10 = 0$$

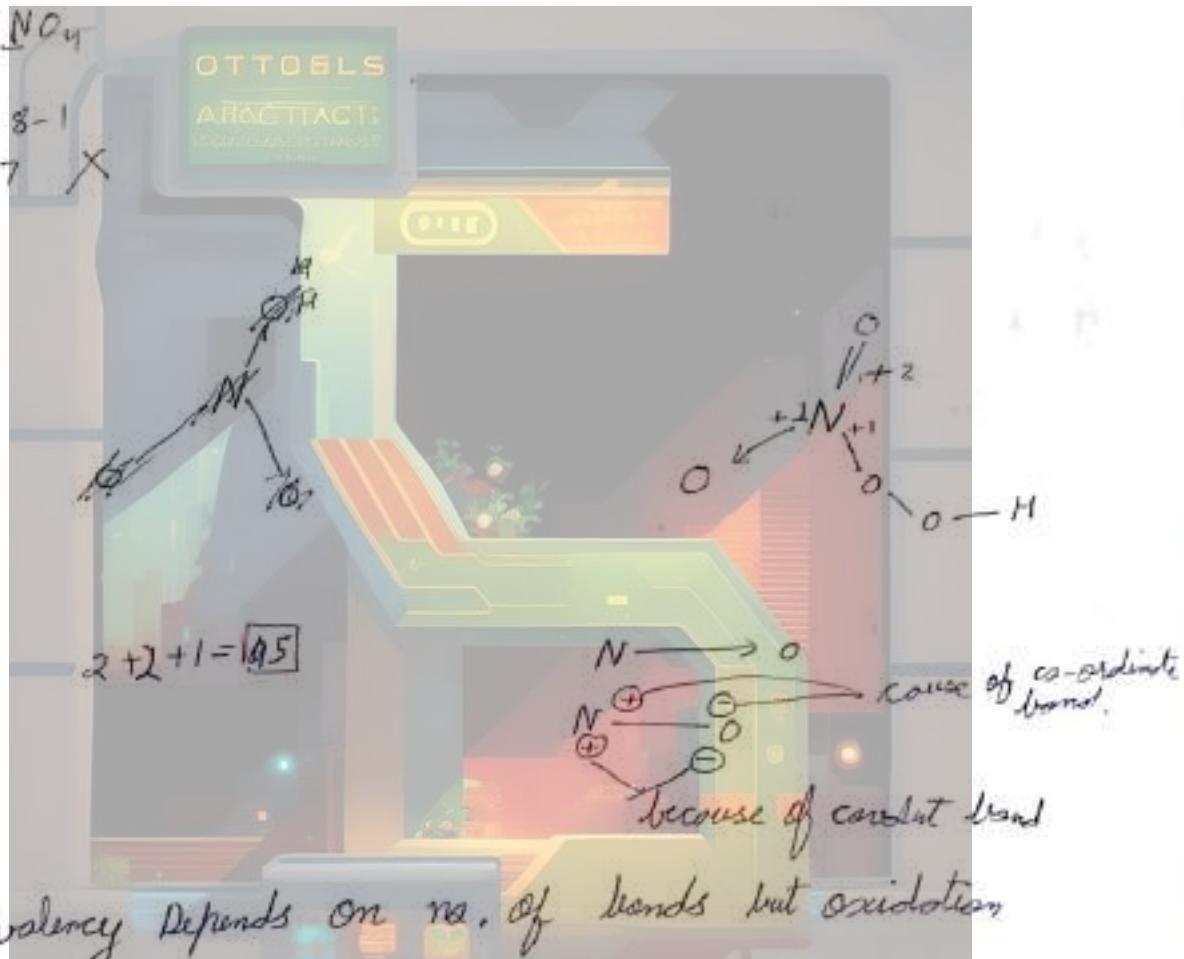
$x = 0 + 10$ but chromium has only 6 valence e^-



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

Eg. NH_4NO_2

$$\begin{aligned} N &= 8 - 1 \\ &= 7 \end{aligned}$$



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

- May or may not equal Δ

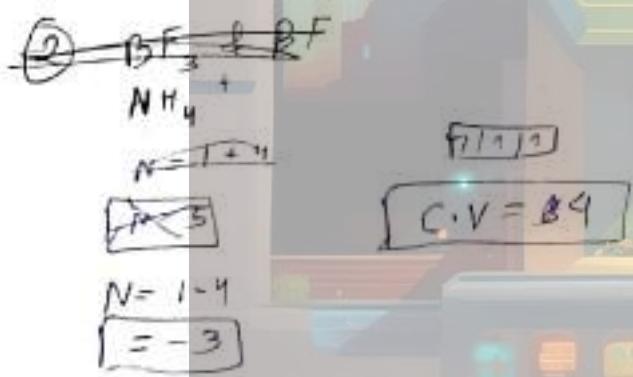
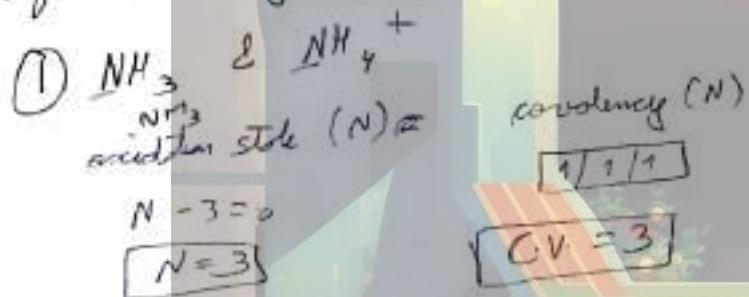
Covalency

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

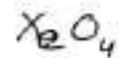
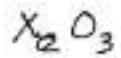
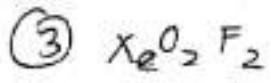
Oxidation state

- ① Permanent formal charge
- ② Can be +ve or fraction.
- ③ Do not depend on no. of bonds.

Q find covalency and oxidation state of



$$\begin{array}{ll} O.S = 3 & O.S = 3 \\ C.V = 3 & C.V = 4 \end{array}$$



$$0.5 = c$$

$$0.5 = 6$$

$$0.5 = 8$$

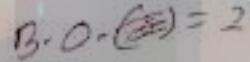
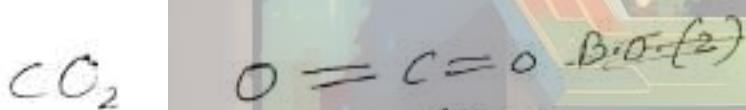
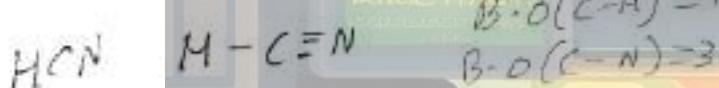
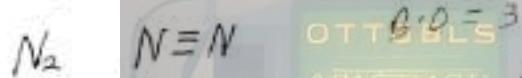
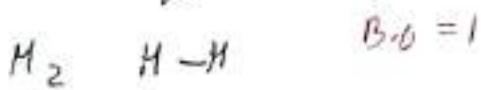
$$C \cdot v = 6$$

$$C \cdot V = 5$$

C.V = 8

Bond Order (B.O):-

→ It is the effective no. of bonds formed between 2 two atoms.

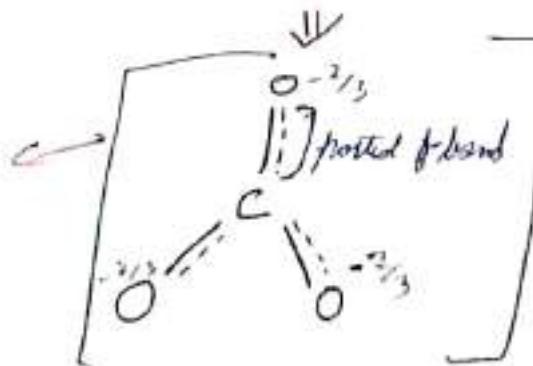


resonating structures or canonical

Resource :-



~~word~~
~~stable~~



→ Resonance Hybrid.

$$0 \cdot 0 = 1 + \frac{1}{3} = \frac{4}{3} \left(= 1 + \lambda \right)$$

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

- 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 stabilizes the molecule as the energy of resonance hybrid is less than energy of any single canonical structure.

conditions for resonance

- ① One $\text{O}=\text{N}-\text{H}$ \Rightarrow double bond
- ② L.P $-=\text{H}$
- ③ $=-\text{H}=\text{H}$
- ④ $\text{O}=\text{N}-\text{H}\equiv\text{H}$

24) $Cr = 6$ ✓

25) 1 +3 ✓

26) -3 ✓

~~27)~~ -3 ✓

28) 3 ✓

29) 5 ✓

30) 16 ✓

31) 1 -3, +5 ✓

32) 2 ✓

33) 3 ✓

34) $\frac{8}{3}$ ✓

35) $\frac{+3}{3}$ ✓

36) $1, 6$ +3, 6 ✓

37) 5 ✓

~~38)~~ 3 ✓

39) 7 ✓

40) 5 ✓

41) 1 ✓

42) -1 ✓

43) 5 ✓

44) 3 ✓

45) 2 ✓

46) 2 ✓

47) 6 ✓

48) -2 ✓

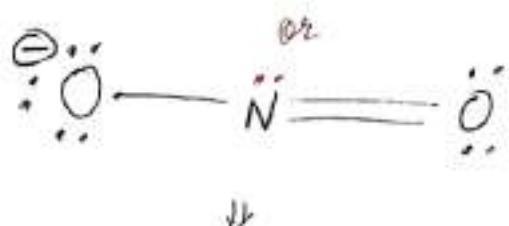
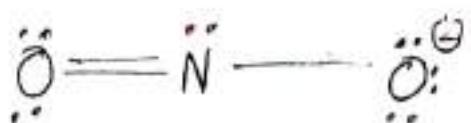
49) -1_3 ✓

50) 0 ✓

51) 2 ✓

OTTOELS
ARCTICACTS
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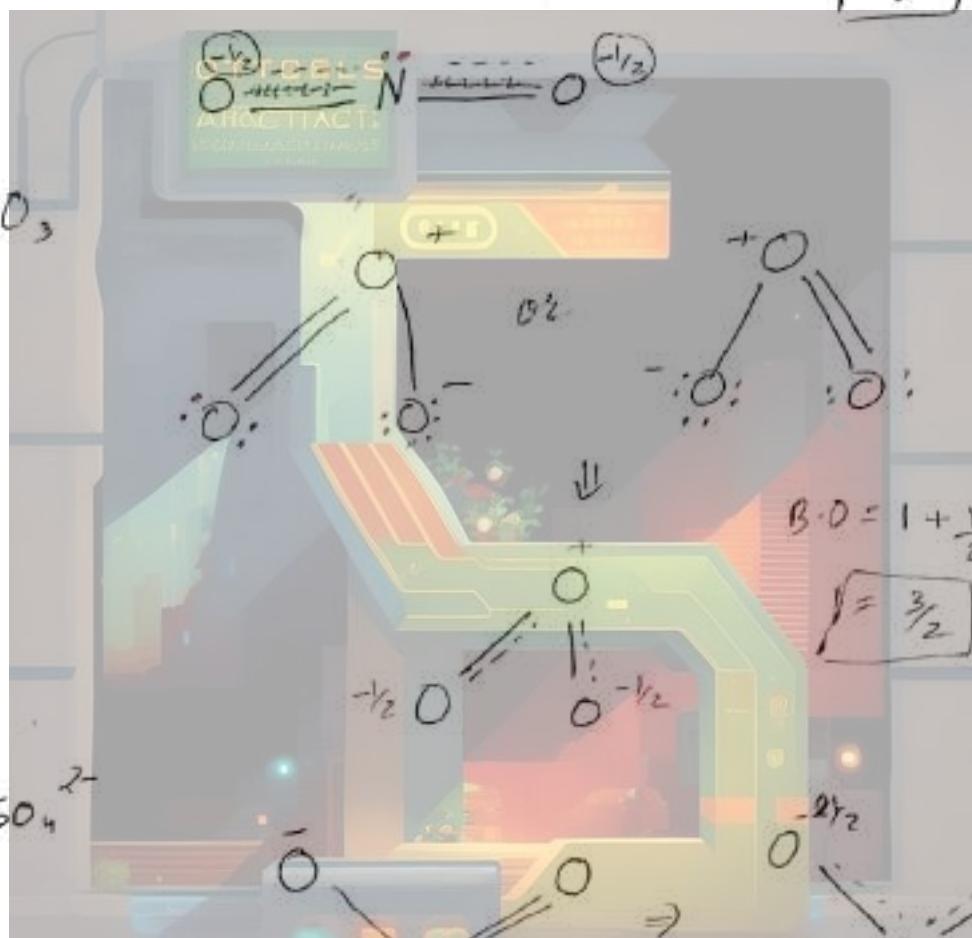
Eg(1) NO_2^-



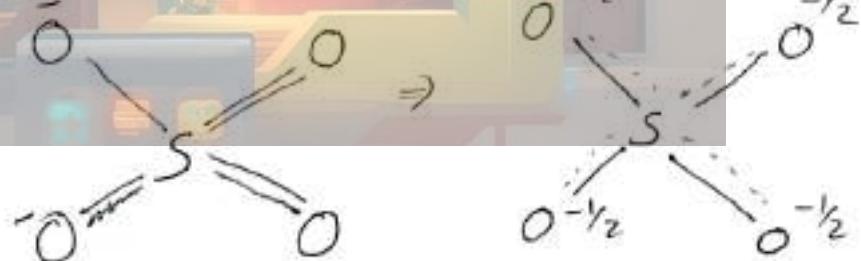
$$B \cdot O = 1 \oplus \frac{1}{2}$$

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

Eg(2)



Eg(3) SO_4^{2-}

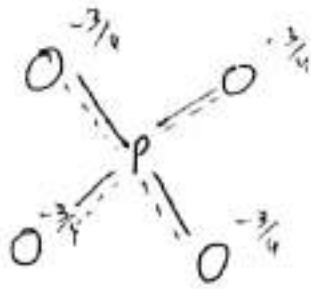


$$= \frac{6}{4}$$

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

(16)

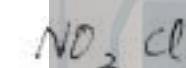
Eg ④ PO_4^{3-}



$$\beta \cdot O = \frac{2(4-1)}{4}$$

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

Eg ③.

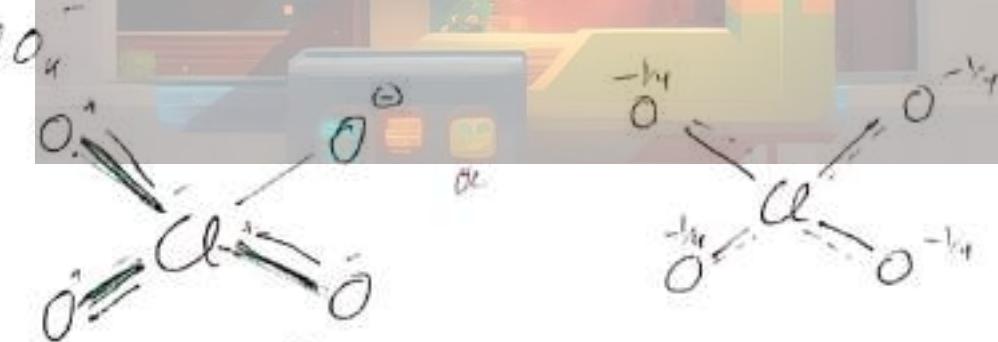


$$\beta \cdot O \text{ } \sigma(N-O) = \boxed{1}$$

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

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

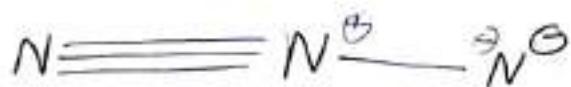
Eg ⑤.



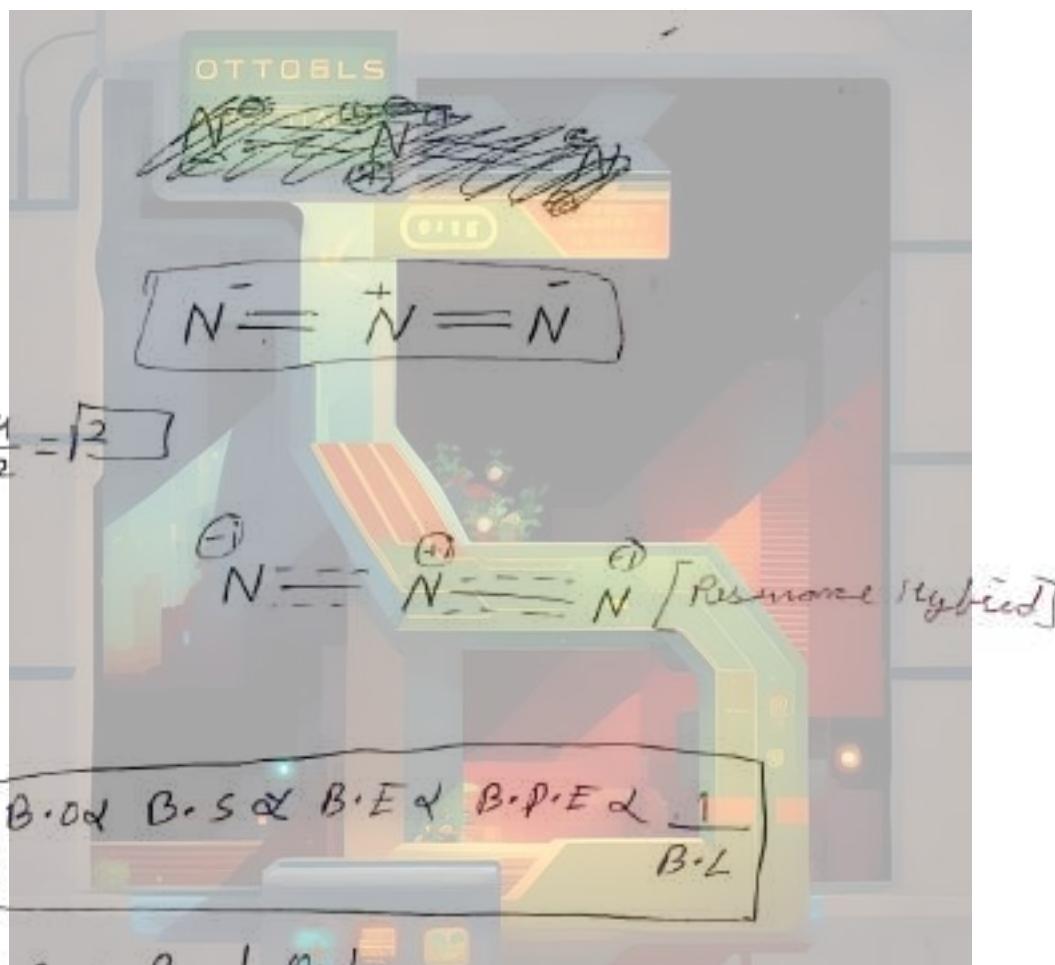
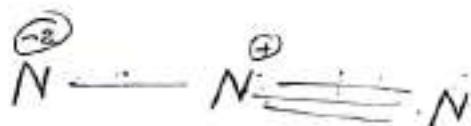
$$\beta \cdot O = \boxed{1} \frac{9+9}{4}$$

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

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



B.O =



Note:-

$$\textcircled{1} \quad B \cdot O \propto B \cdot S \propto B \cdot E \propto B \cdot P \cdot E \propto \frac{1}{B \cdot L}$$

B · O → Bond Order

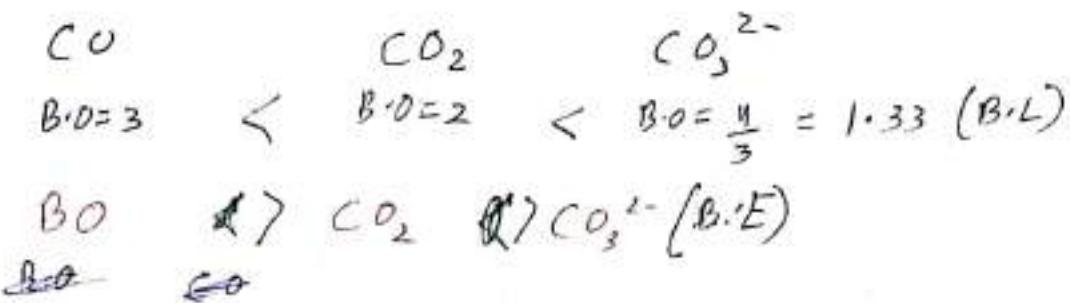
B · S → Bond Strength

B · E → Bond Energy

B · D · E → Bond Dissociation energy

B · L → Bond Length

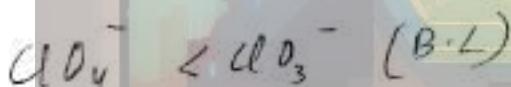
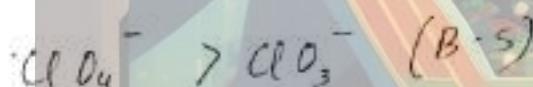
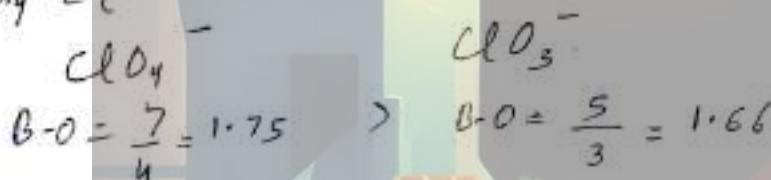
Q Compare bond length in (O-C-O)



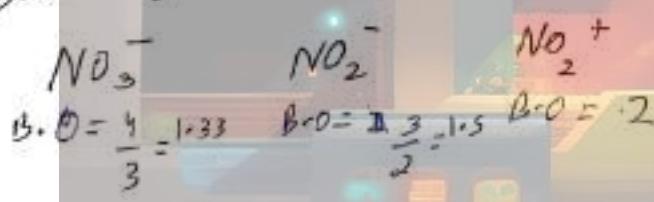
Q Compare (S-O) Bond Lengths in



Q Compare (Cl-O) Bond strength & r Length

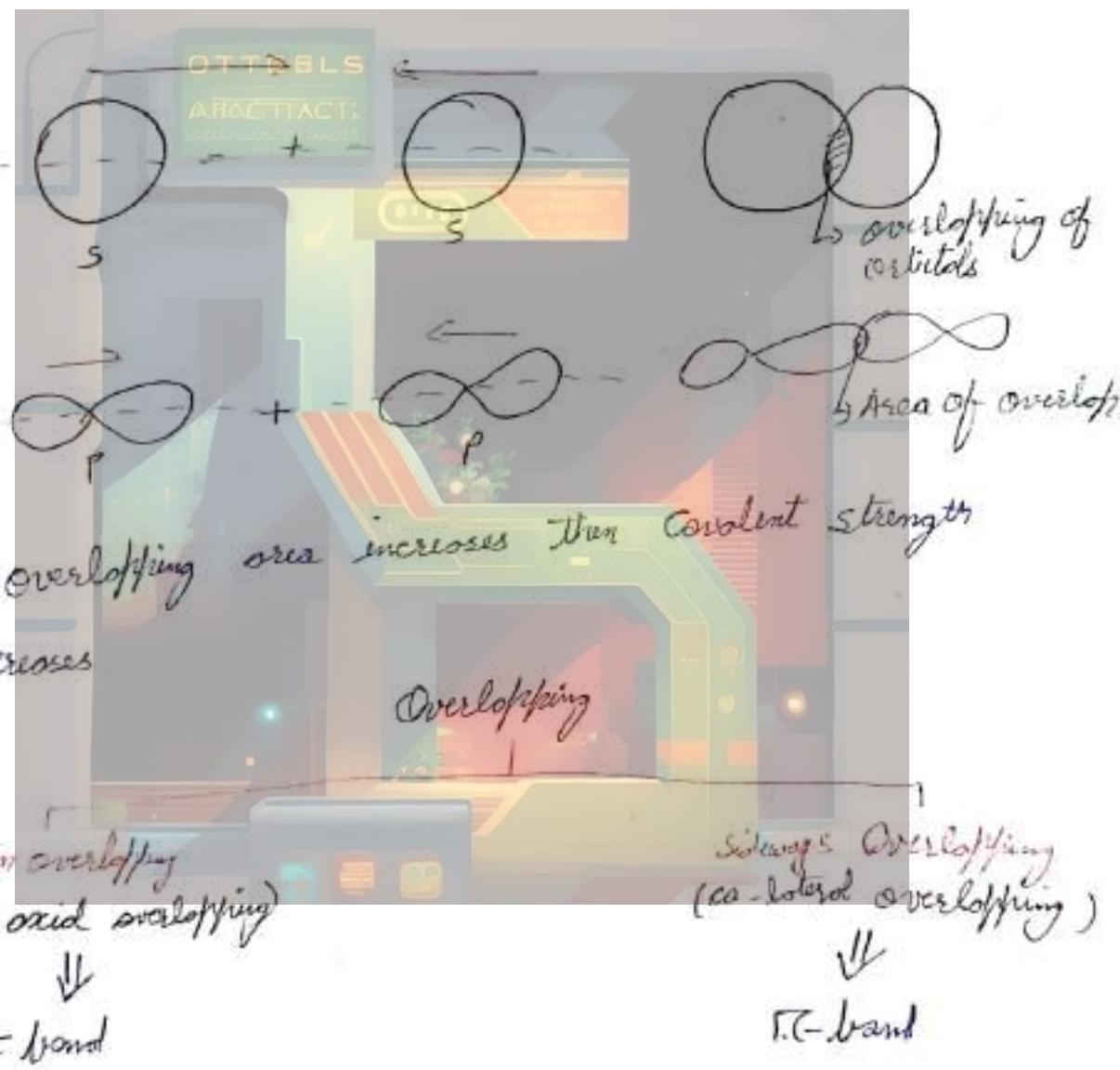


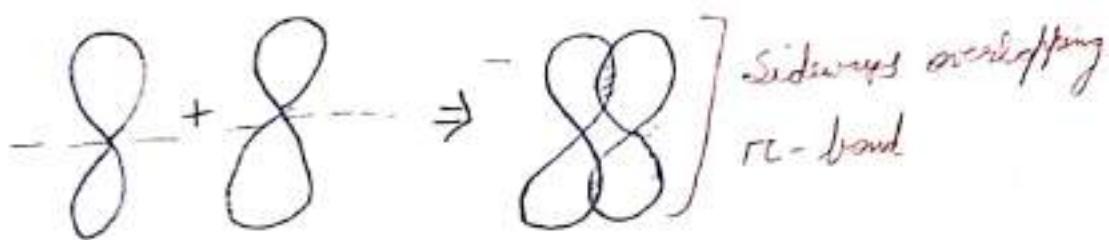
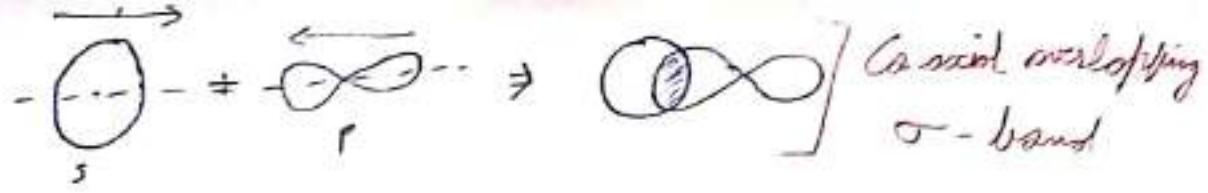
Q (N-O) bond length



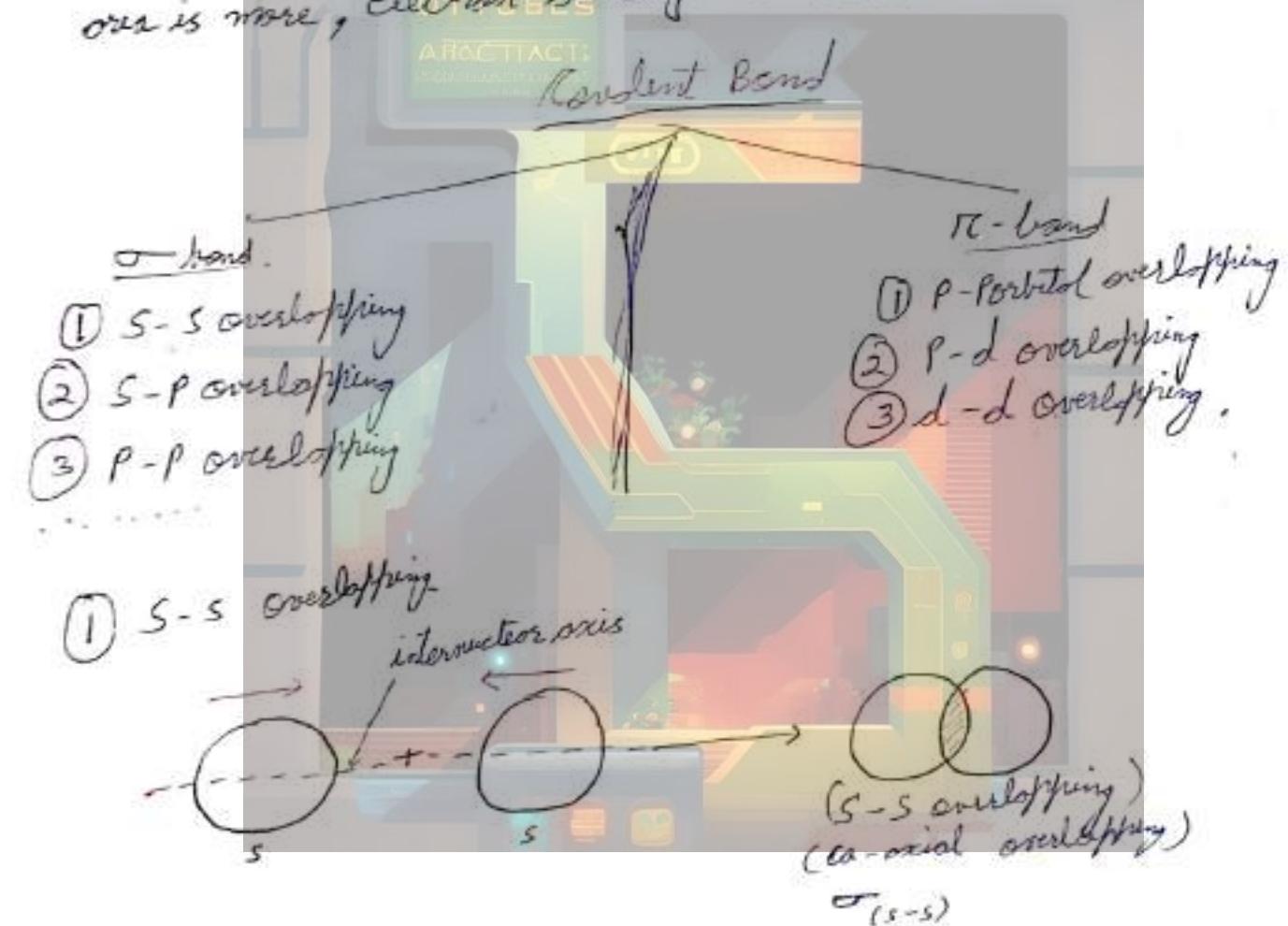
Valence Bond Theory (VBT)

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

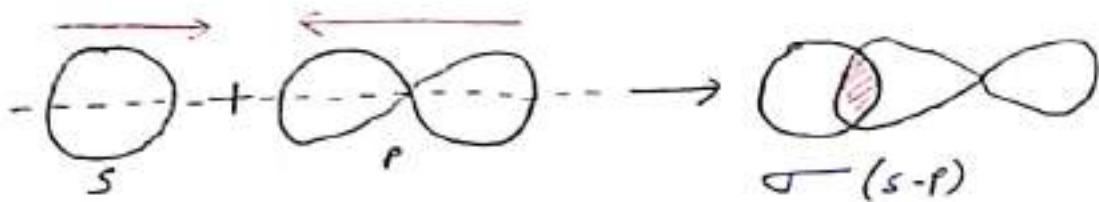




Note \rightarrow σ bond is stronger than π bond as overlapping area is more, electron density is more.



② $s - s$ Overlapping



③ $p - p$ Overlapping



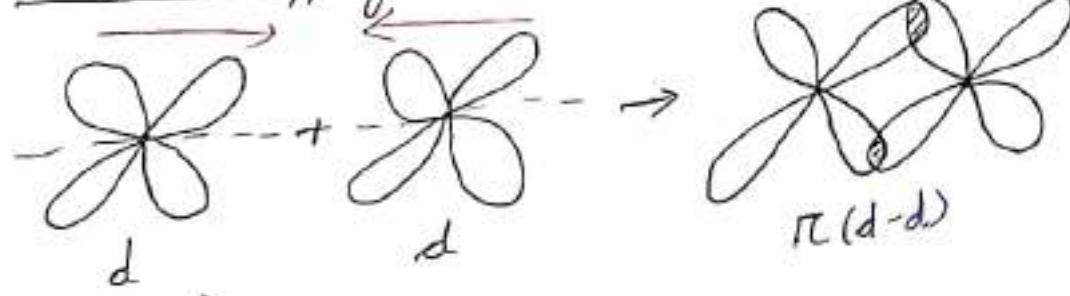
④ $p - p$ Overlapping



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 interlocked known as delta bond (δ).

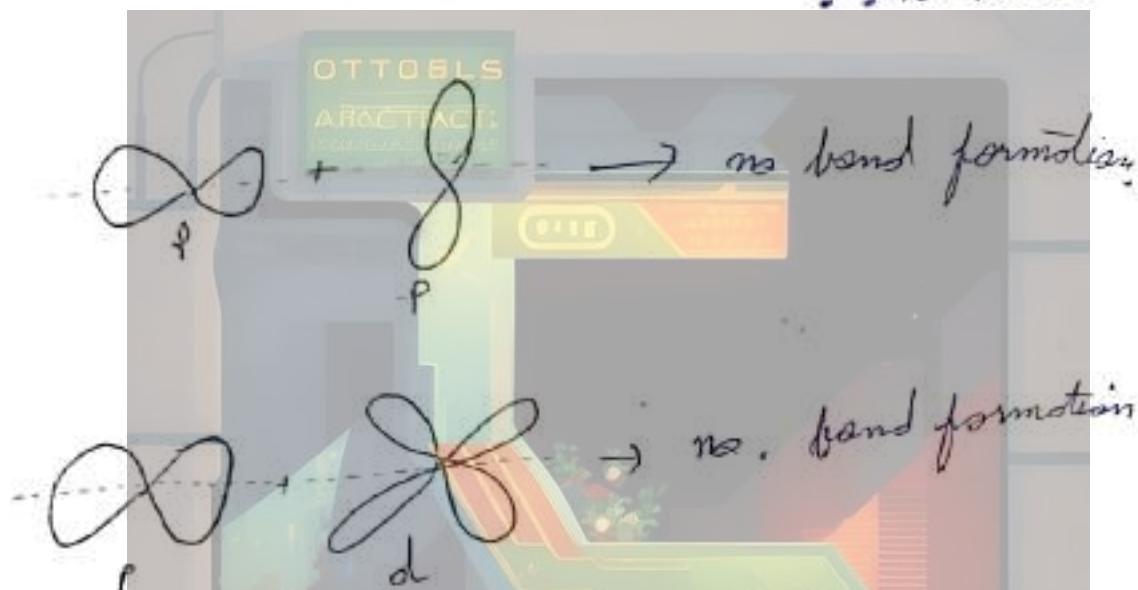
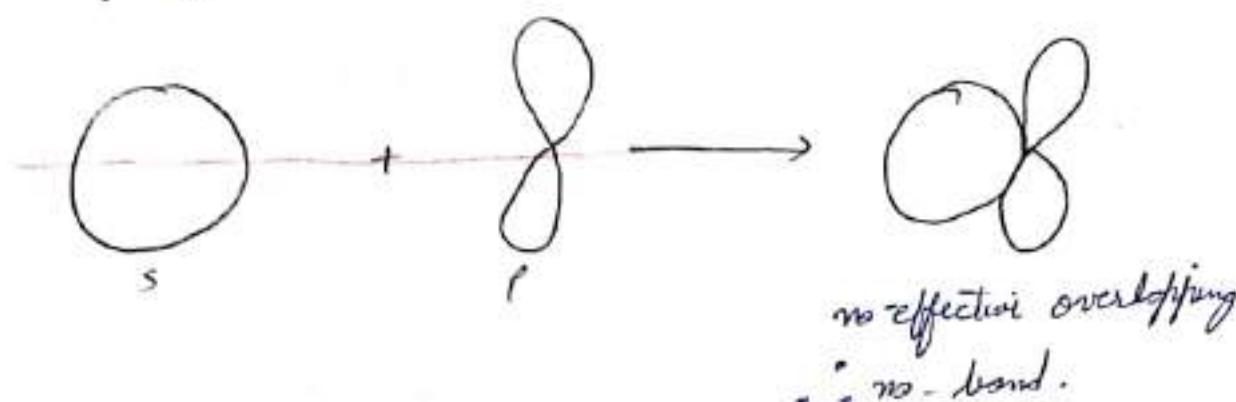
Note:- w.r.t one orbital

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

Type of bond
 σ bond

π bond
 δ 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, d-orbital & never form σ -bond.
- ③ The first bond formed between two atoms is always a σ -bond.
- ④ Once a σ bond is formed, the next bonds formed after that will be π -bond.

Atomic orbitals

Interactions 197-25

Type of bond

① $s + p_x$

$x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$

σ
 x
 x

② $s + p_y$

$x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$

x
 σ
 x

③ $s + p_z$



x
 x
 σ

④ $p_x + p_z$

$x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$

σ
 π
 π

⑤ $p_{ay} + p_y$

$x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$

π
 σ
 π

⑥ $p_z + p_z$

$p_x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$

π
 π
 σ

⑦ $p_c + p_y$

$x \longrightarrow$
 $y \longrightarrow$
 $z \longrightarrow$

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
 π
 x

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

$x \rightarrow$

OTTO
YLS
ARCTIC
VOLKSWAGEN

$x \rightarrow$

$y \rightarrow$

$z \rightarrow$

x
 x

π

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

$x \rightarrow$

$y \rightarrow$

$z \rightarrow$

π

π

δ

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

$x \rightarrow$

$y \rightarrow$

$z \rightarrow$

δ
 π

π

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

$x \rightarrow$

$y \rightarrow$

$z \rightarrow$

π

δ

π

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

$x \rightarrow$

$y \rightarrow$

$z \rightarrow$

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)

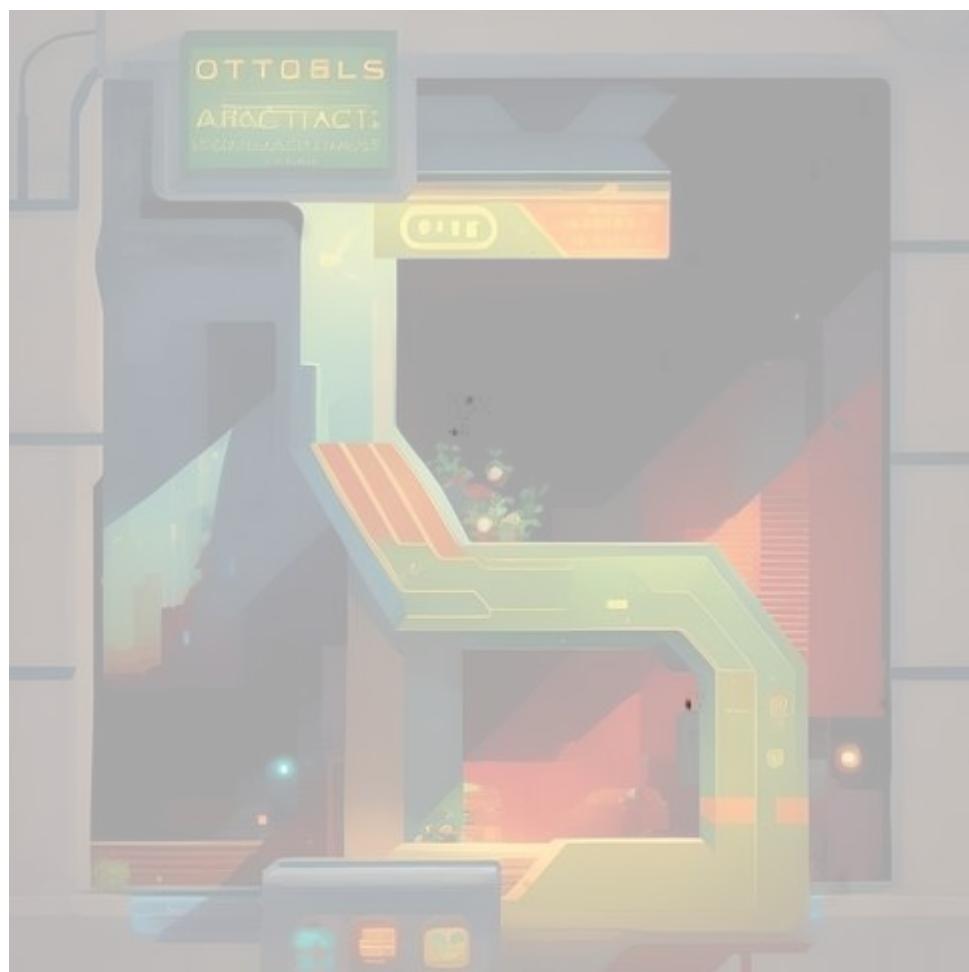
H.W. (29 -05-2024)

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

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

S-1 $\{16\}$

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

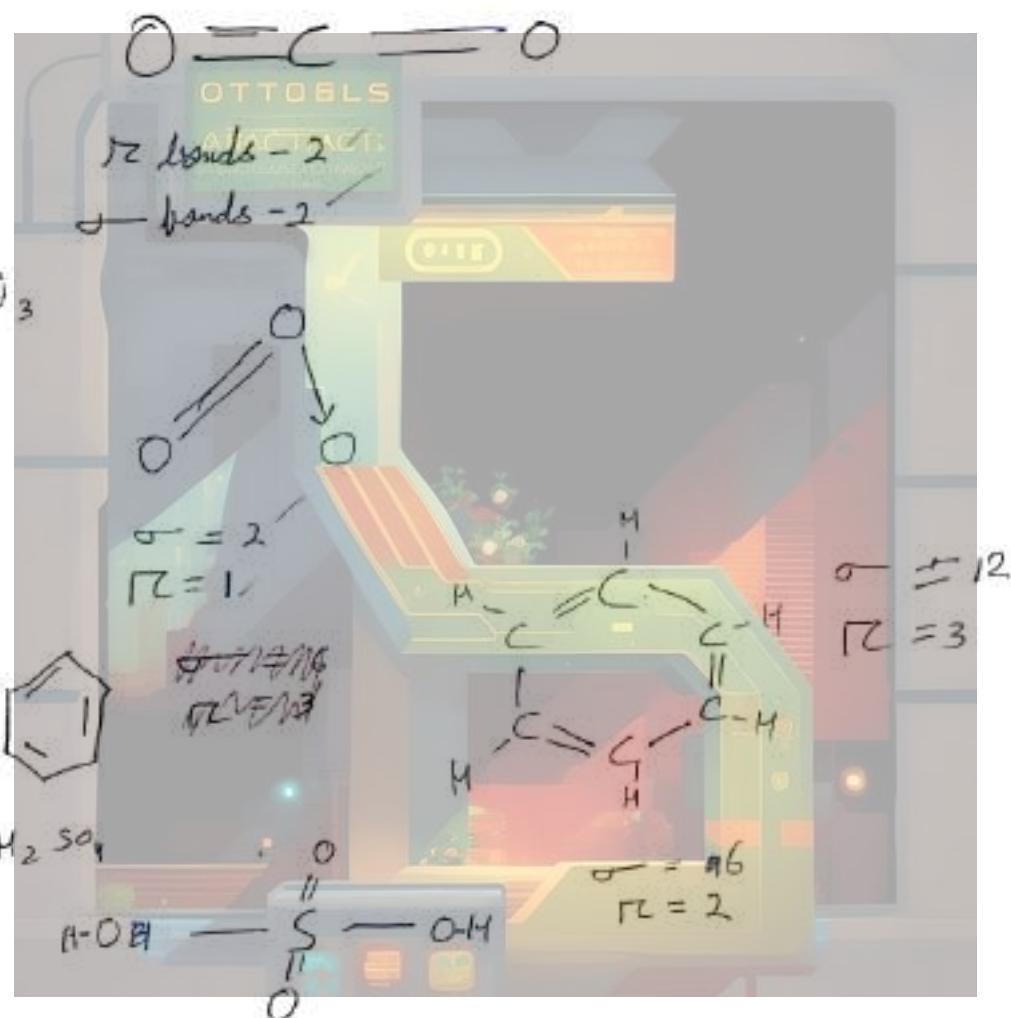


Note -

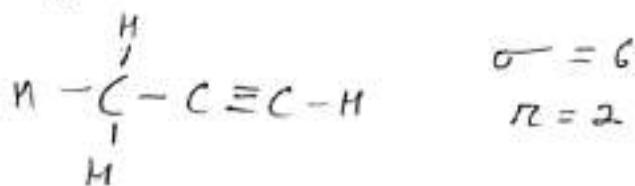
- ① Two perpendicular orbitals can never form any bond.
- ② π bond is formed only after ~~sigma~~ σ bond.
- ③

Q Calculate π bond & σ bond?

① CO_2

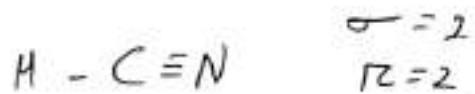


⑤ $\text{H}_3\text{C}-\text{C}\equiv\text{C}-\text{H}$



(28)

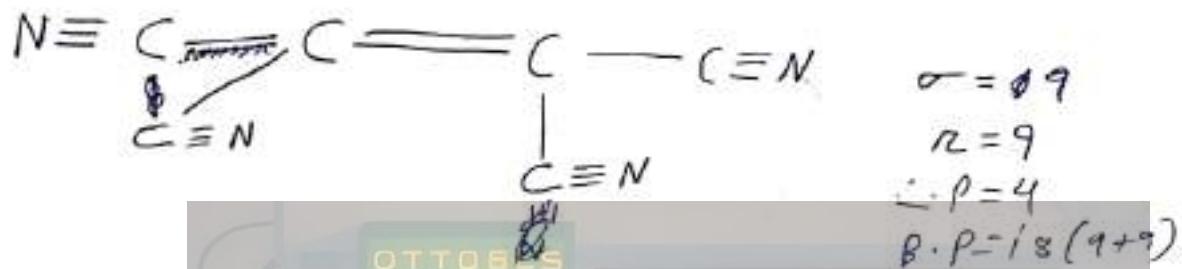
⑥ HCN



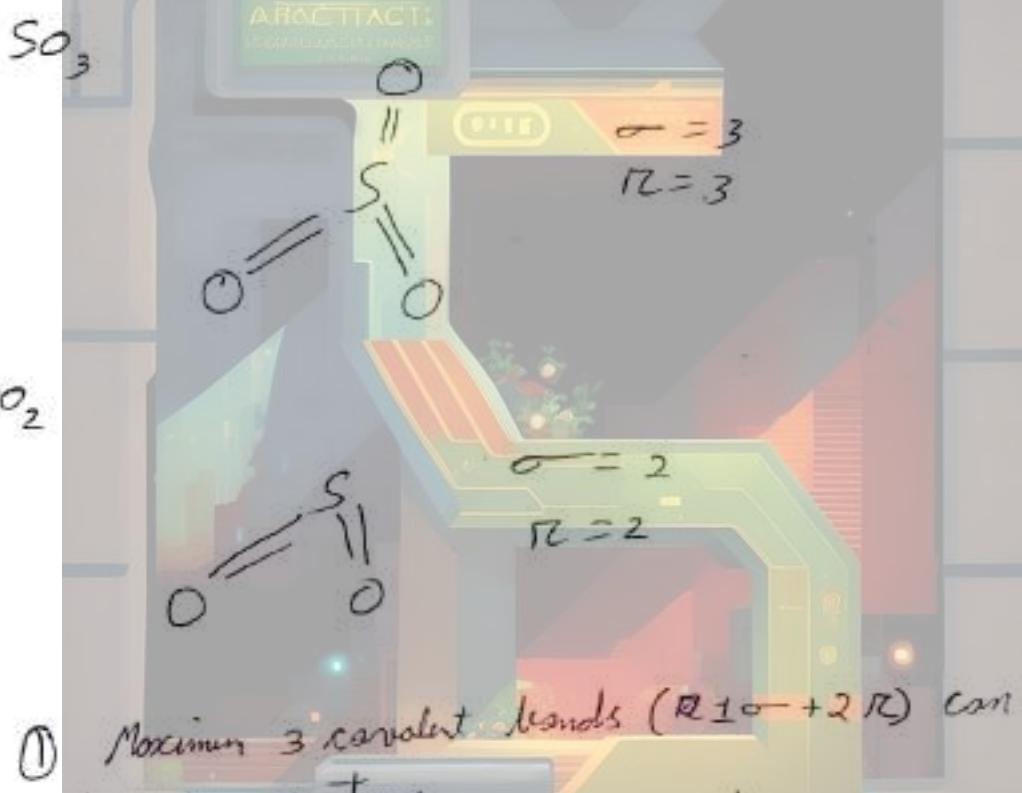
$$\sigma = 2$$

$$r = 2$$

⑦ $\text{C}_2(\text{CN})_4$



⑧ SO_3



⑨ SO_2

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

Eg $\text{N} \equiv \text{N}$ $\text{C} \equiv \text{O}^+$ $\text{C} \equiv \text{F}$

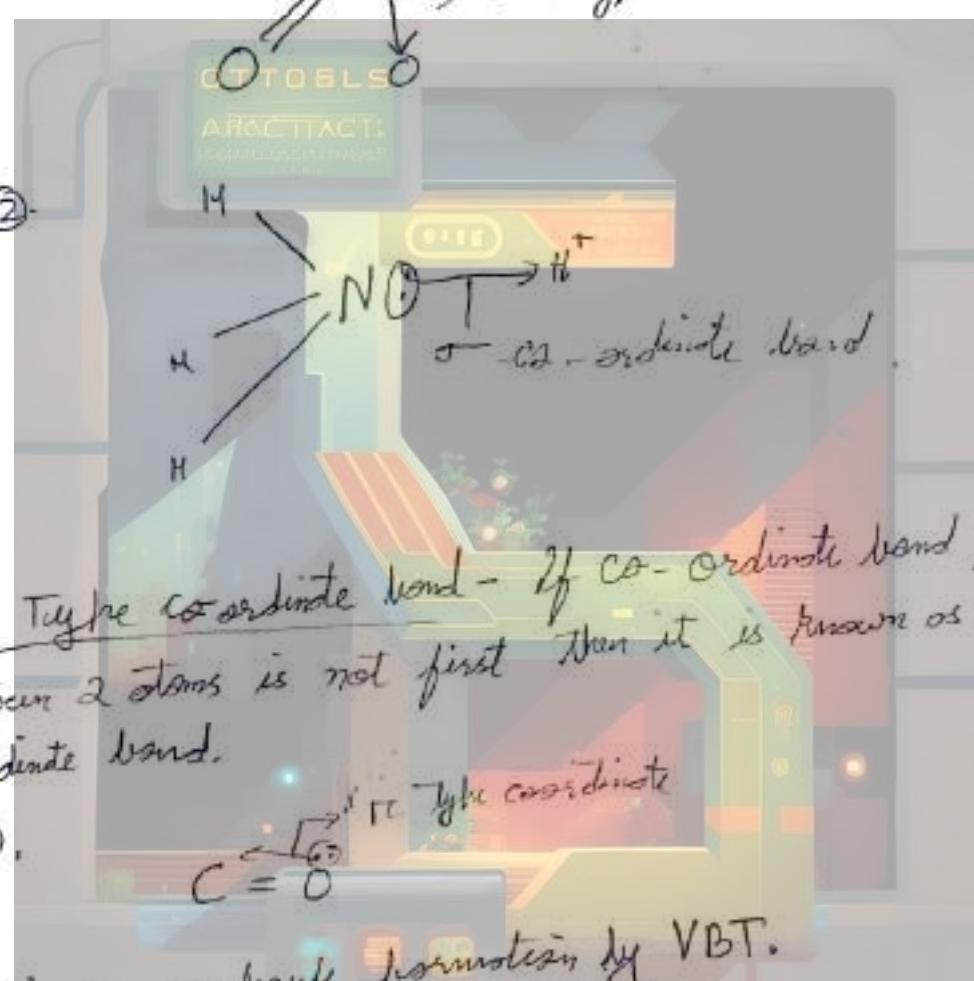
Types of Co-ordinate bonds

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

e.g.

O₂

e.g(2)

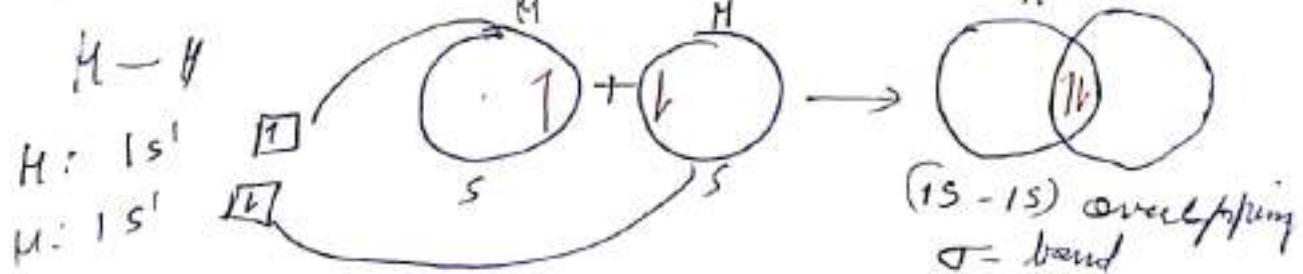


- ② π Type co-ordinate bond - If co-ordinate bond formed between 2 atoms is not first then it is known as π type co-ordinate bond.

e.g(1).



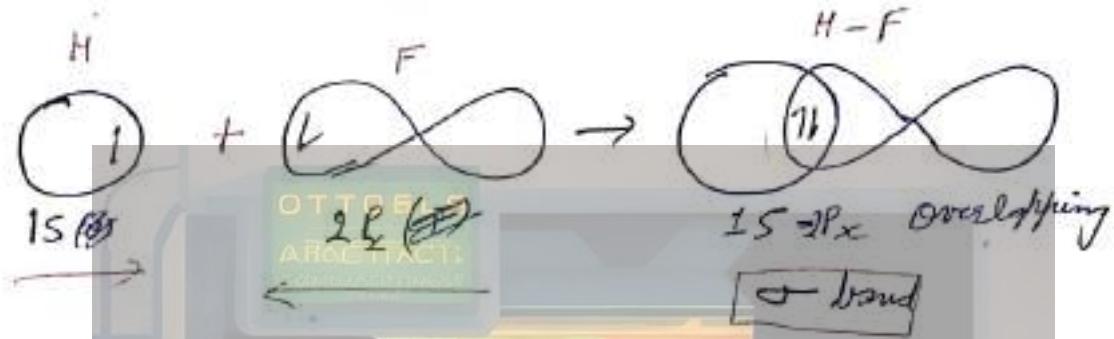
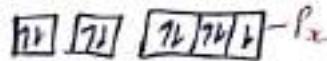
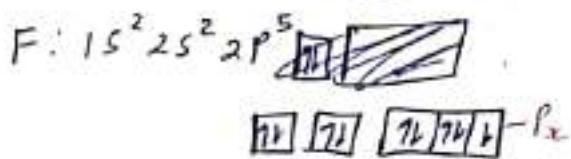
- Q Explain H₂ molecule formation by VBT.



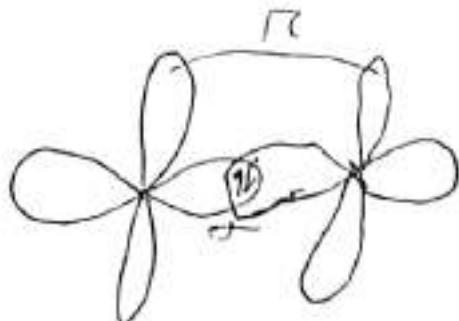
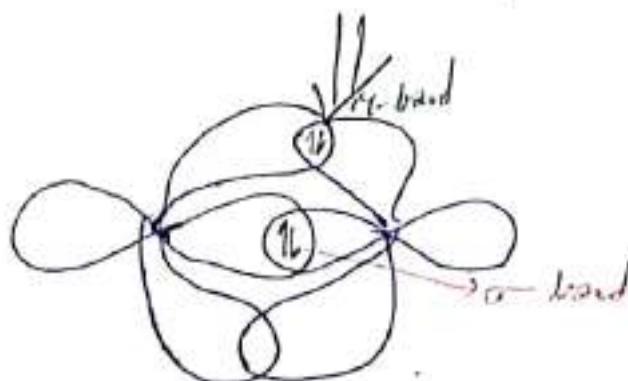
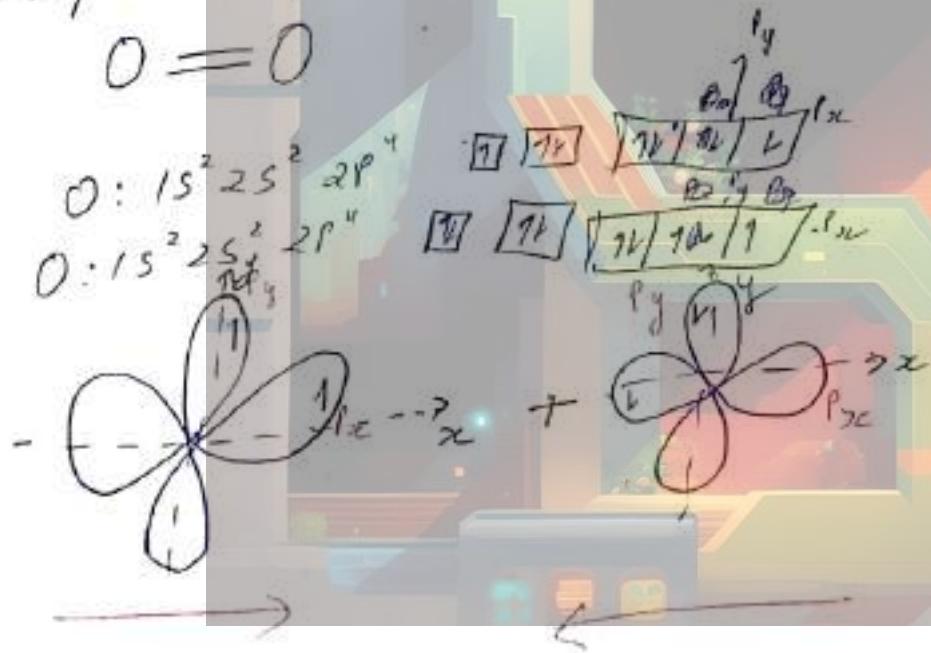
Q formation of HF explaining VBT



1



Q Explain O₂ formation using VBT.



(3)

Q explain N_2 formation by VBT?

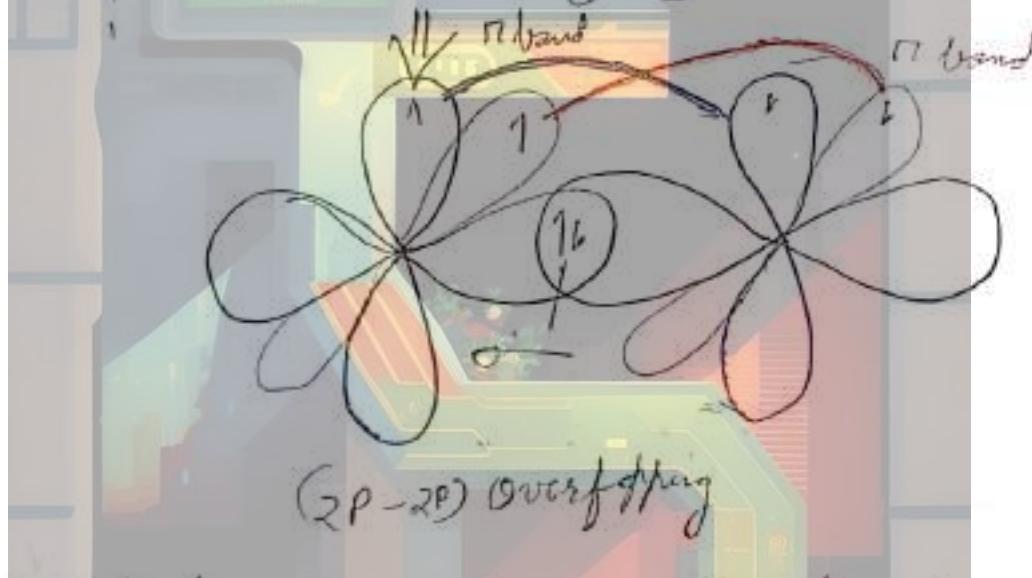
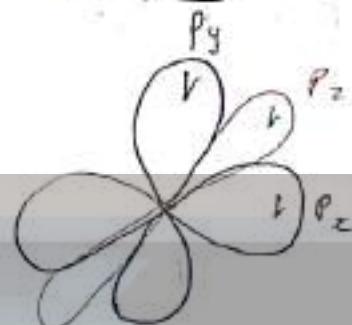
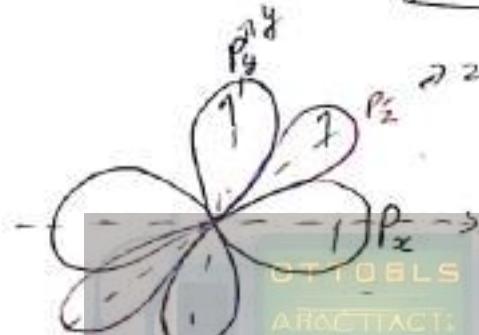
$N \equiv N$

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

$1s$	$2s$	$2p_x$	$2p_y$	$2p_z$
1	1	1	1	1

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

$1s$	$2s$	$2p_x$	$2p_y$	$2p_z$
1	1	1	1	1



σ -bond

π -bond

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

- ①② Co-axial 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 no.)

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

$$(1s-1s) > (2s-2s) > (3s-3s) \text{ over bond strength}$$

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

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

Case 2 when n is same
Bond Strength & directional nature of orbital.

directional nature $\Rightarrow s, p < d < f$
 n decides over directional nature

Q. Compare bond for σ -bond

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

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

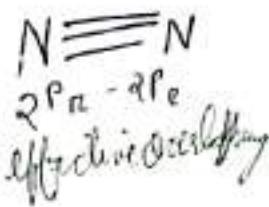
④ $2P_{N_2} - 3P_{N_2}$ $2P_{N_2} - 3d_{N_2}$ $2P_{N_2} - 4P_{N_2}$

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

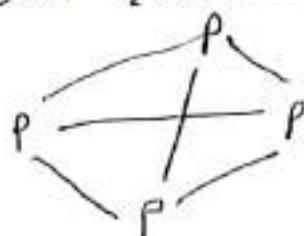
⑤ $2P_{N_2} - 2P_{N_2}$ $2P_{N_2} - 3d_{N_2}$ $2P_{N_2} - 3P_{N_2}$ $3P_{N_2} - 3P_{N_2}$

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

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



$P \equiv P$
 $2P_{N_2} - 2P_{N_2}$
not effective so not effective
instead exist as

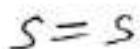


Q At Room Temperature O_2 exist but S_2 does not why.



$$2P_n - 2P_n$$

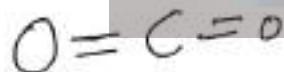
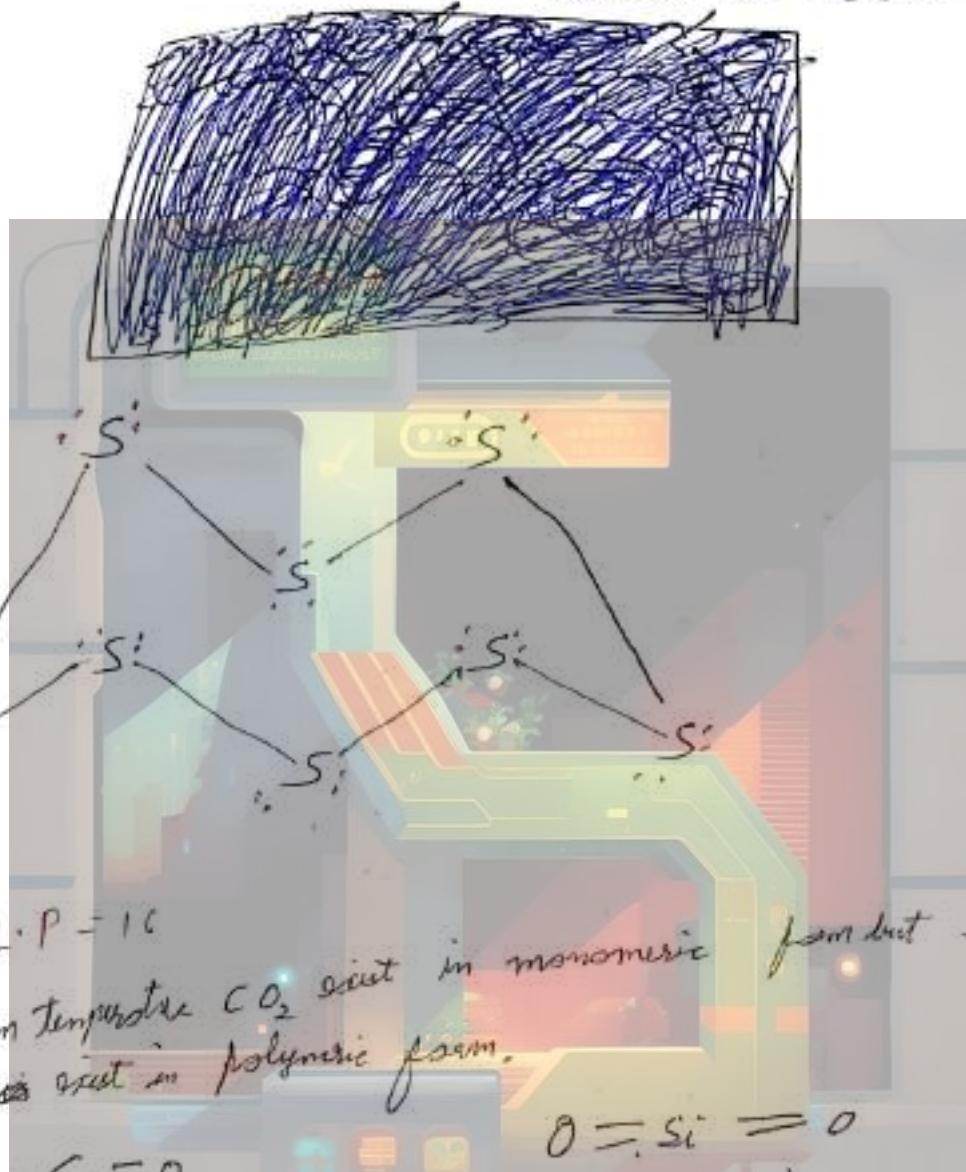
effective overlapping



$$3P_n - 3P_n$$

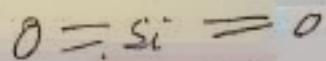
not effective

so exist as S_8 (crystalline structure)



$$(2P_n - 2P_n)$$

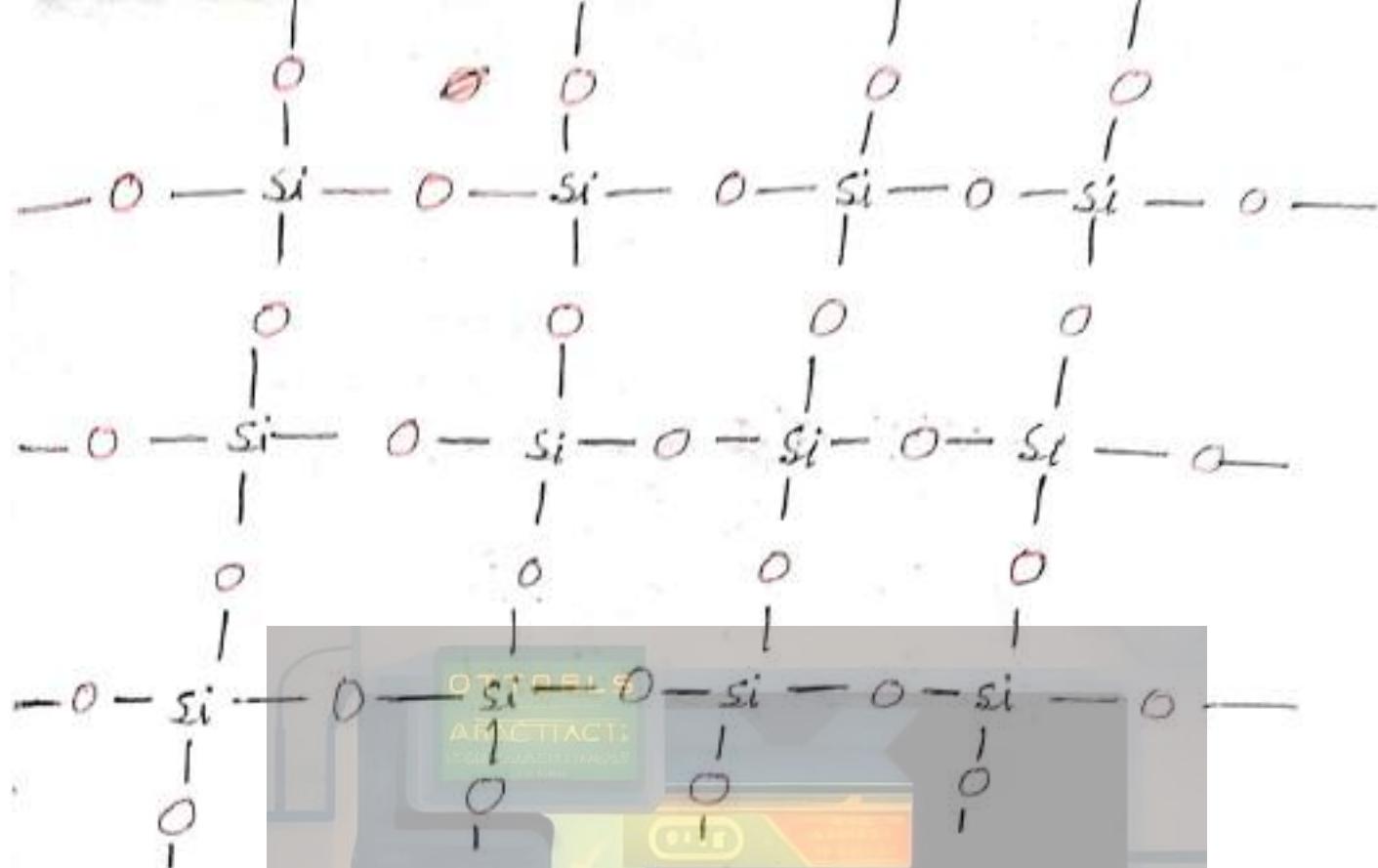
effective



$$(2P_n - 3P_n)$$

unstable

so exist in polymeric form
 $(SiO_2)_n$



Bond Strength Order

Single Bond

- ① $-c - c > -s_i - s_i$
 - ② $-n - n < -p - p$
 - ③ $-o - o < -s - s$
 - ④ $f - f < c - c$

Multiple Bond

- ① $-c=c \rightarrow -\underset{1}{c}=\underset{1}{c}$
 - ② $\ddot{n}=n \rightarrow \ddot{\underset{1}{n}}=\underset{1}{n}$
 - ③ $\ddot{o}=o \rightarrow \ddot{\underset{1}{o}}=\underset{1}{o}$

due to small size LIP
repulsion dominate over n

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



expected: $\text{F}_2 > \text{Cl}_2 > \text{Br}_2 > \text{I}_2$

Real: $\text{Cl}_2 > \text{Br}_2 > \text{F}_2 > \text{I}_2$

Q Bond strength



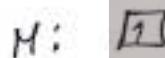
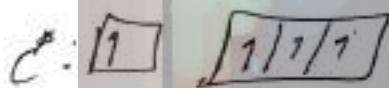
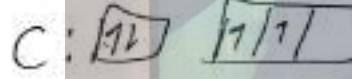
(1s-2s) (1s-2p) OTTO (1s-4p) (1s-5p)

ABSTRACTS
VOLUME 10 NUMBER 1

Drawbacks of VBT

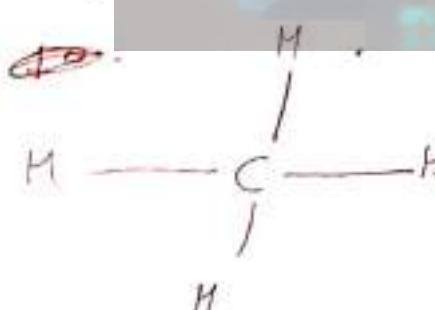
① Valence Bond Theory fails to explain geometry and bonding of molecules containing more than 2 atoms.

① CH_4



\downarrow σ -bond: 1s-2s (weaker)

\downarrow σ -bond: 1s-2p (stronger)



→ According to VBT CH_4 molecule has all

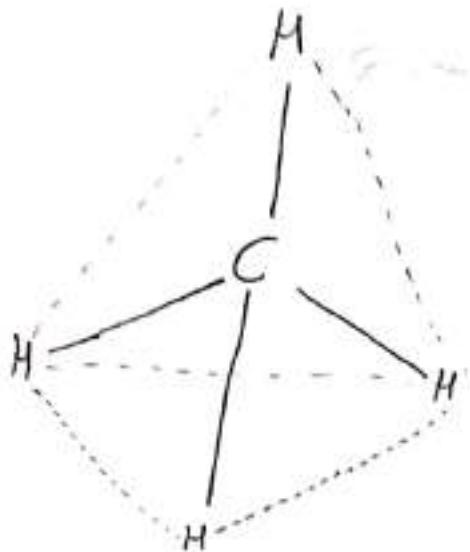
\downarrow (C-H) bond length & bond strength equal

→ 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 (hybridization).

1. bond is weak so it should have longer bond length \Rightarrow
but in reality all the 4 bonds have ~~all~~ some bond length.
have

according to VBT

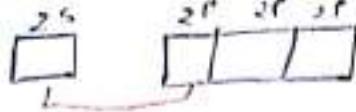


Hybridisation

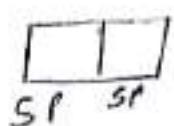
→ It is the best molecular concept to explain geometry of molecules.

Assumptions -

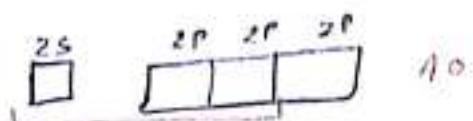
- ① Hybridisation is the process of intermixing of atomic orbitals which has comparable energy or less energy difference.
- ② As many no. of atomic orbitals are participated, some no. of hybrid orbitals will be formed.
- ③ Once the hybrid orbitals are formed they arrange themselves in such a way so that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



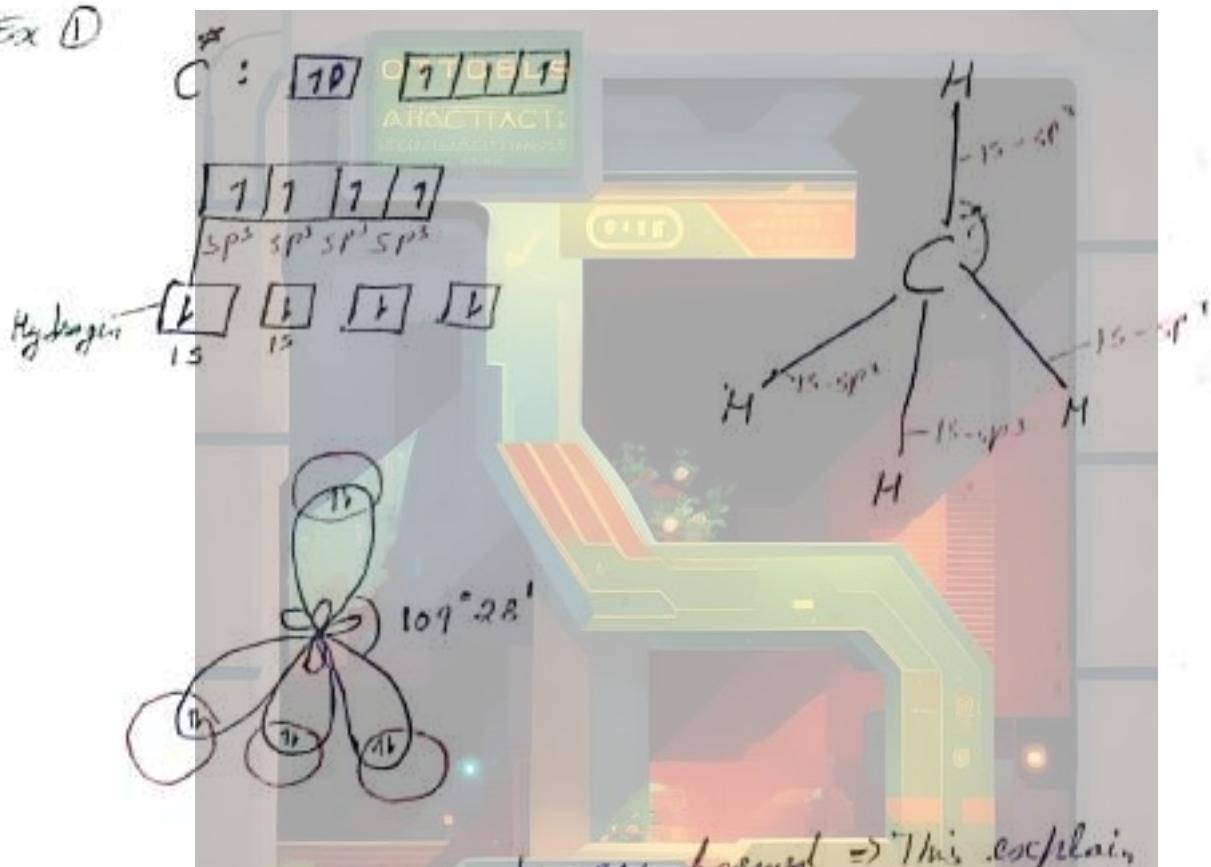
Atomic orbitals (AO)



Mixed orbital (HO)

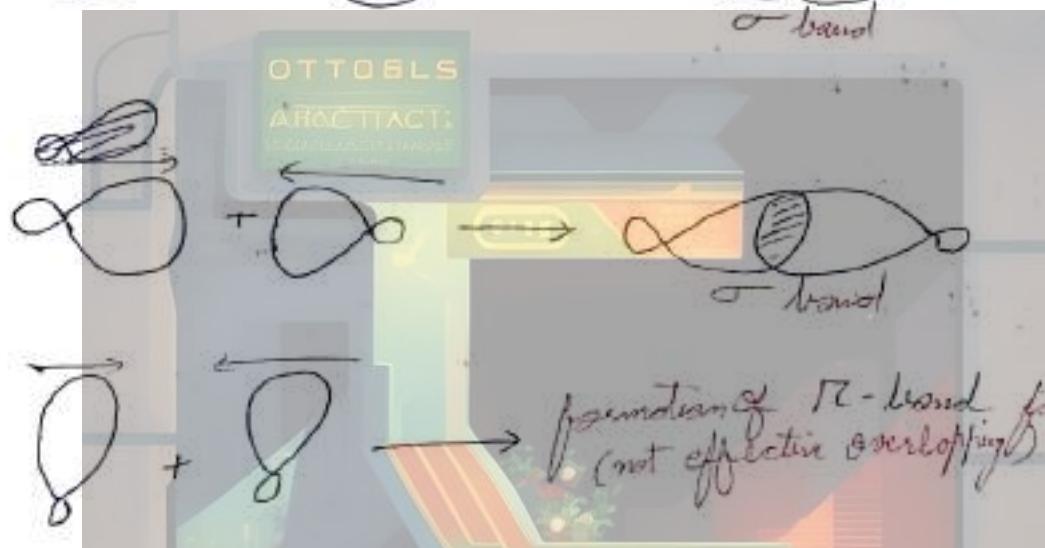
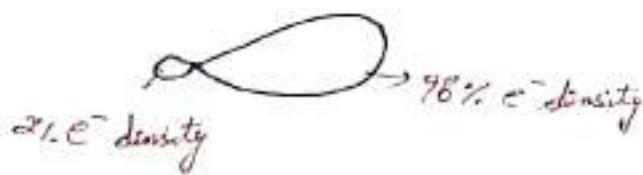


Ex ①



- 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)

Shape of Hybrid Orbital



Formed hybrid orbitals have some % S character
% P character
% d character
length & strength

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

$SP < SP^2 < SP^3 \}$ length of H-O \propto % s-character $\propto \frac{1}{\% s\text{-character}}$

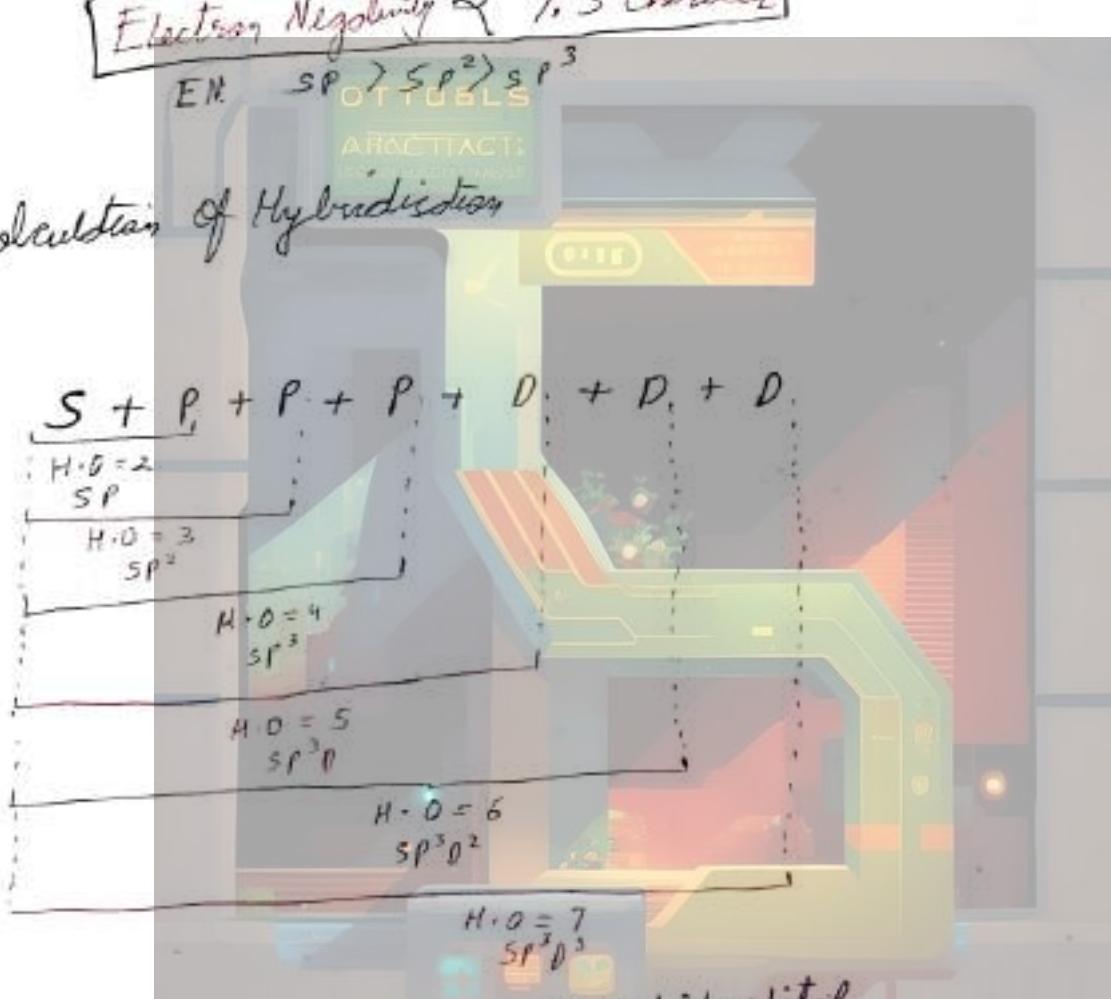
$SP < SP^2 < SP^3 \}$ Energy of H-O \propto $\frac{\% s\text{-character}}{\% p^2\text{-character}}$

$SP > SP^2 > SP^3 \}$ stability of H-O \propto % s-character

Electron Negativity \propto % s-character

EN. $SP > SP^2 > SP^3$

Calculation of Hybridisation



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

\Rightarrow $\#$ no. of σ -bonds + no. of L.P. on central atom
of central atom

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

SN-

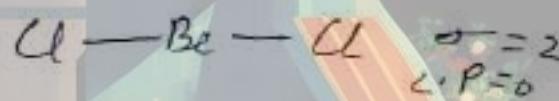
Type of Hybridisation

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

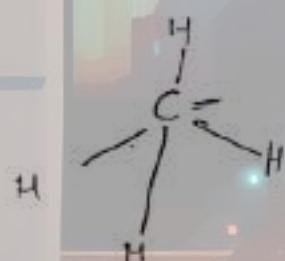
7

OTTOELS
ABSTRACTS
INFORMATION SP^3D^3

(Q) Find the hybridisation of central atom?

(1) $BeCl_2$ 

$$\begin{aligned} SN. &= \sigma + L \cdot P \\ &= 2 + 0 = 2 (sr) \checkmark \end{aligned}$$

(2) CH_4 

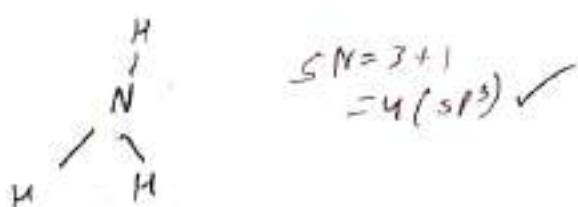
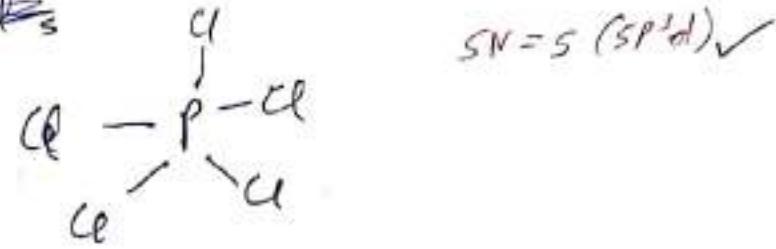
$$\begin{aligned} \sigma &= 8 \cdot 4 \\ L \cdot r &= 0 \end{aligned}$$

$$S.N. = 8 + 0 = 8 (sp^3)$$

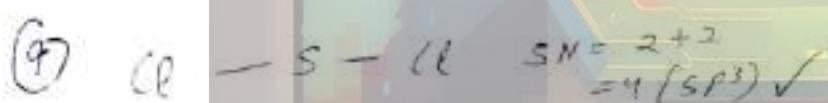
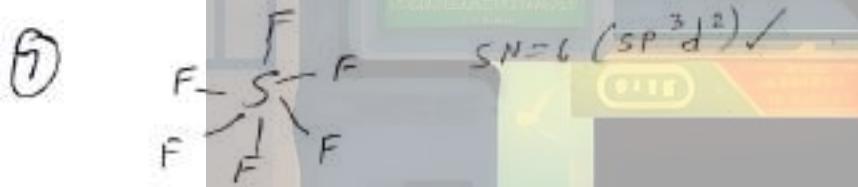
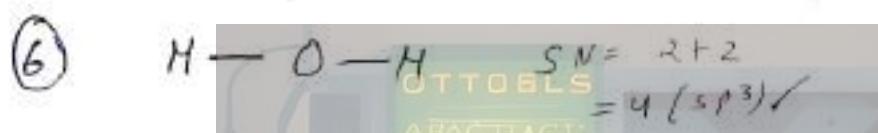
(3) $AlCl_3$

$$\begin{aligned} S.N. &= 3 + 3 + 0 \\ &= 3 (sp^3) \checkmark \end{aligned}$$

- ④ PCl_5 ④ ~~BF_3~~ ⑤
 ⑤ NH_3 ⑥ H_2O ⑦
 ⑥ SF_6 ⑧ IF_7 ⑨
 ⑦ SCl_2 ⑩ XeF_2 ⑪
 ⑧ XeF_4 ⑪ XeF_6

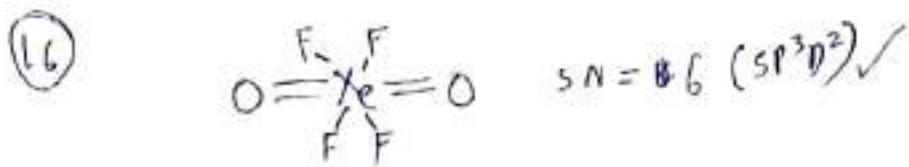
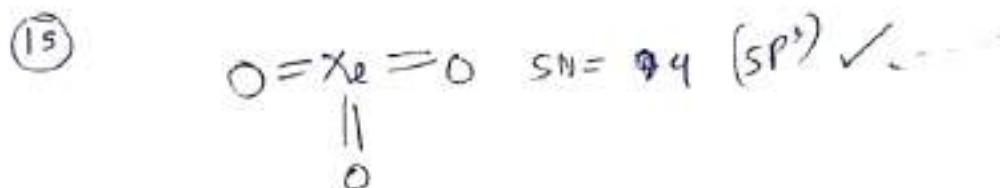
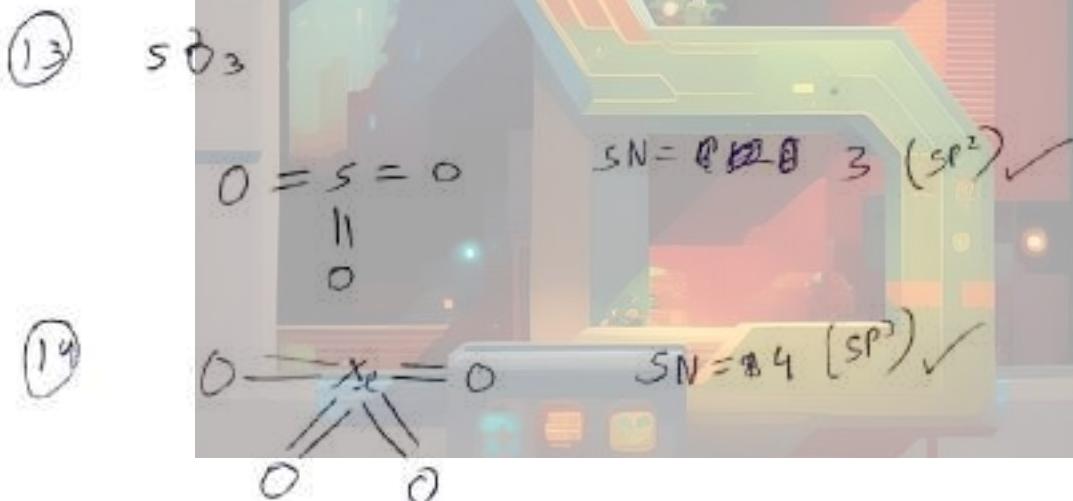
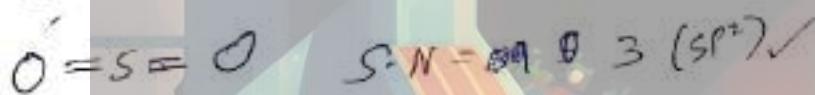


$$SN = 3 + 1 = 4 (\text{sp}^3) \checkmark$$

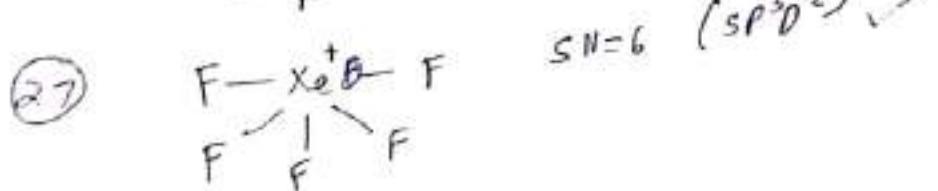
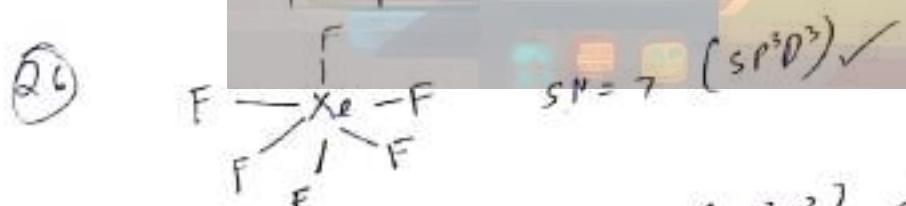
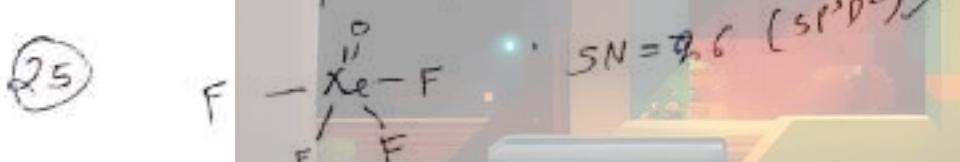
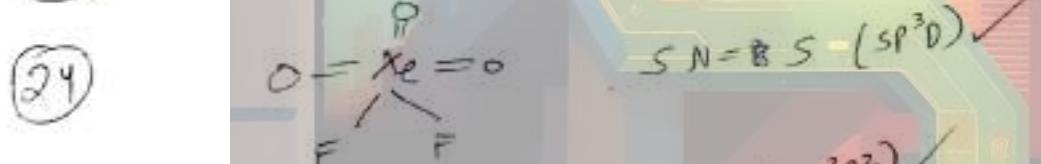
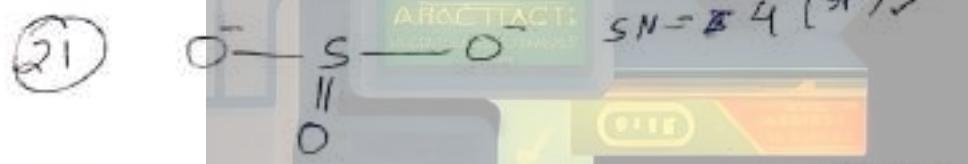
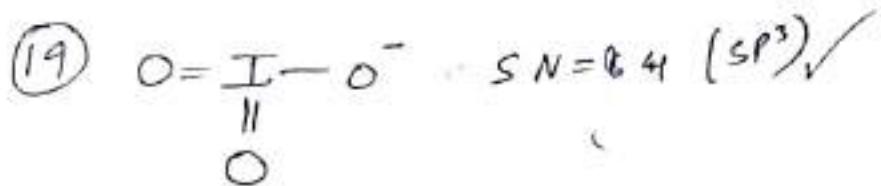


- (12) SO_2 (21) SO_3^{2-} (30) XeF_5^- (49) ClO_4^{-1}
 (13) SO_3 (22) C_2H_4 (31) IF_{47}^- (41) H_3O^+
 (14) XeD_4 (23) C_2H_2 (32) IF_c^- (42) SF_4
 (15) XeO_3 (24) XeO_3F_2 (33) I_3^+ (43) SeF_4
 (16) XeO_2F_4 (25) $XeOF_4$ (34) C_2H_6 (44) $TcCl_4$
 (17) CO_2 (26) XeF_6 (35) BF_2 (45) ClF_3
 (18) $PoCl_3$ (27) XeF_5^+ (36) SO_2SO_3 (46) ICl_2^-
 (19) IO_3^- (28) NO_2^+ (37) O_3^- (47) BeF_3
 (20) $SnCl_3^-$ (29) NO_2^- (38) NH_4^+ (48) BrF_2^-
 (49) XeO_6^{4-} (50) BeF_4^-

- (51) $IeCl_6$ (52) ICl_4^- (53) IF_7 (54) IF_6^- (55) XeF_6
 (56) XeO_4



(44)



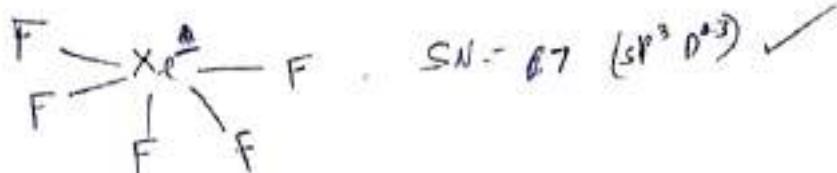
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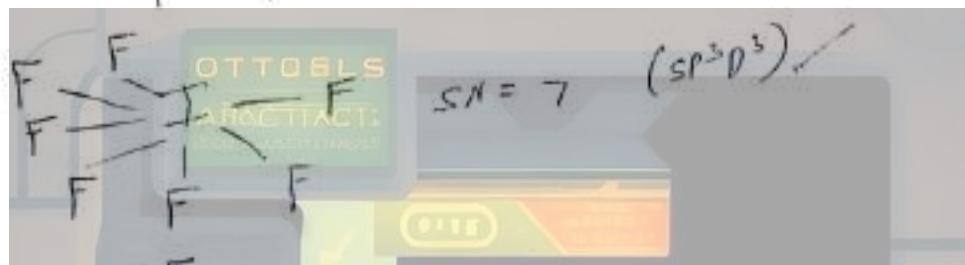
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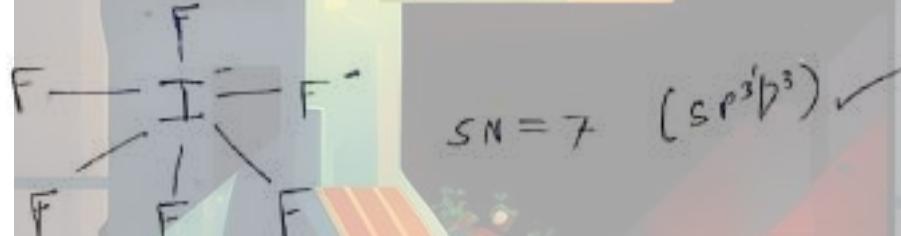
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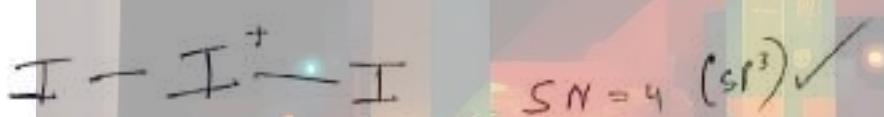
(31)



(32)



(33)



(34)



(34.5)

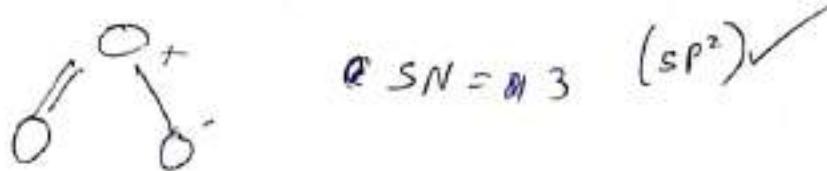


(46)

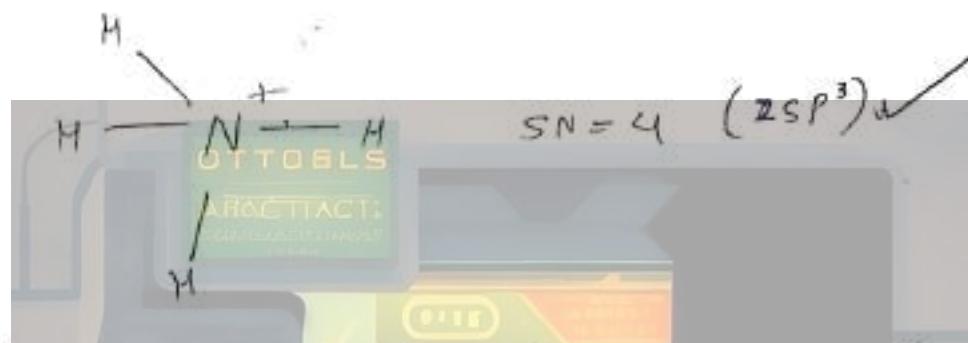
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(37)



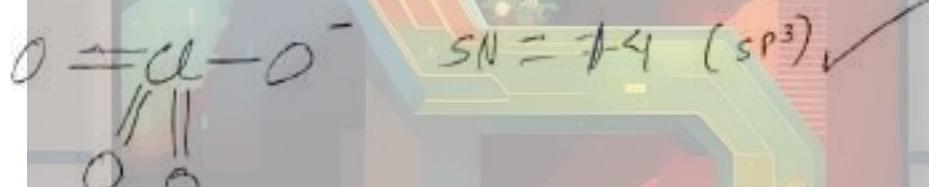
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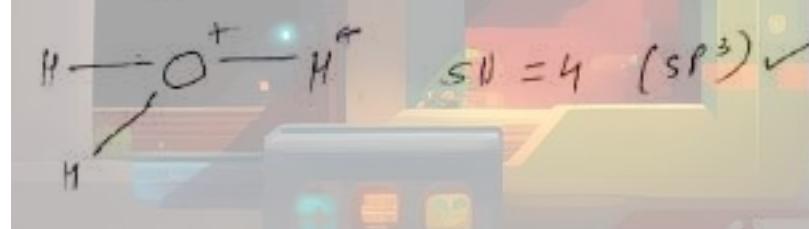
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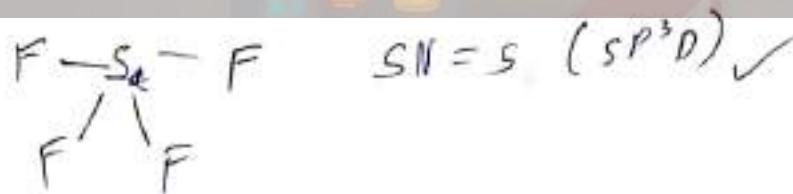
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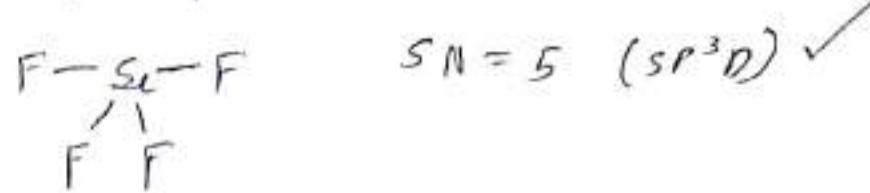
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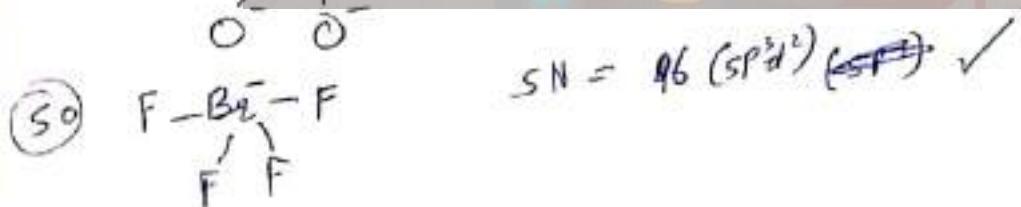
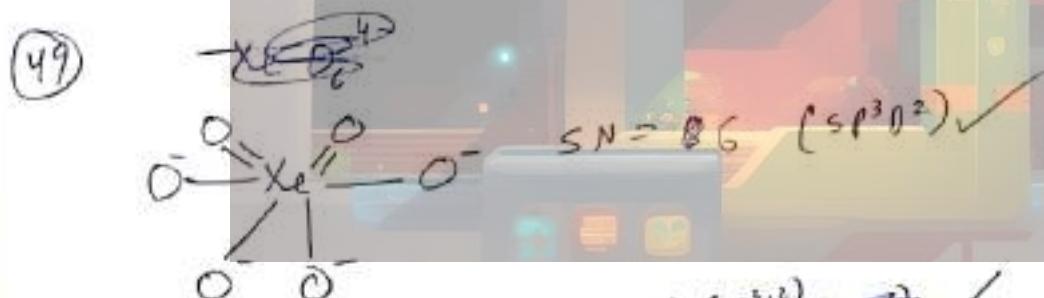
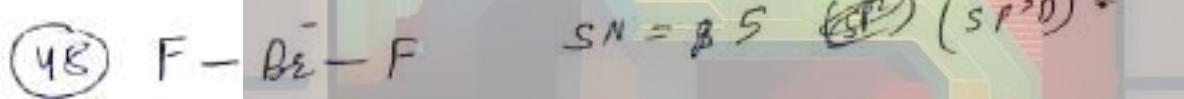
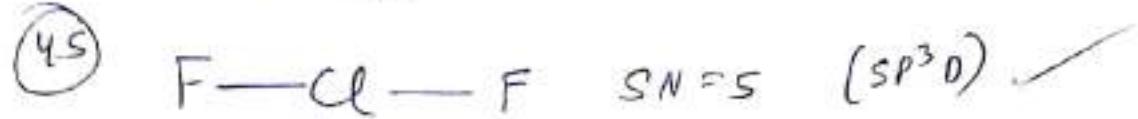
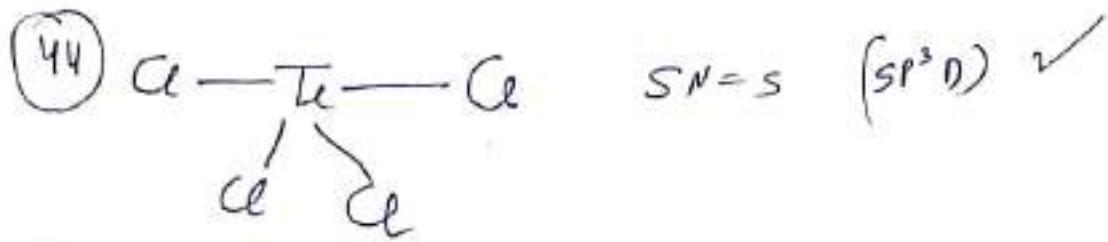
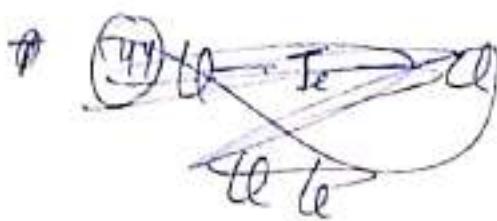
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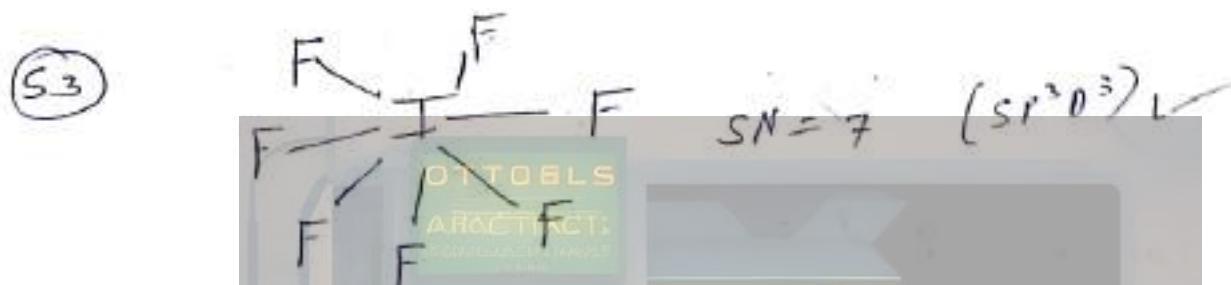
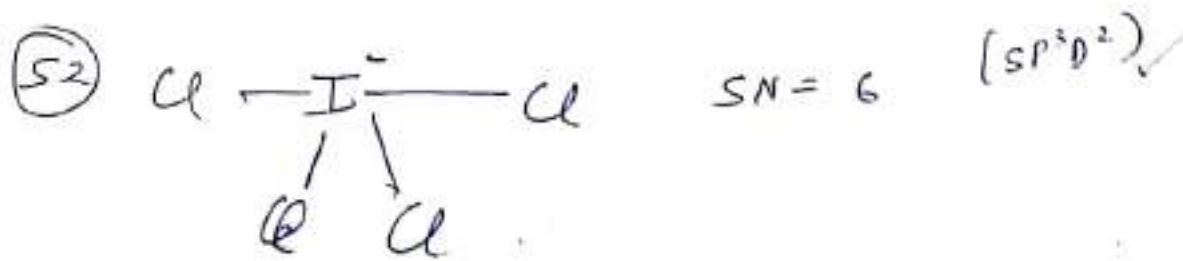
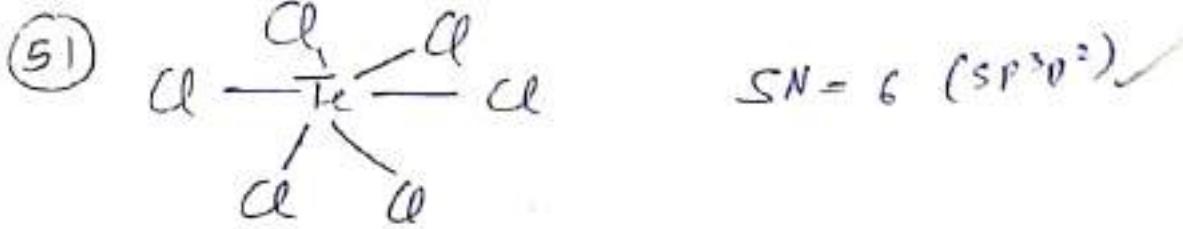
(43)



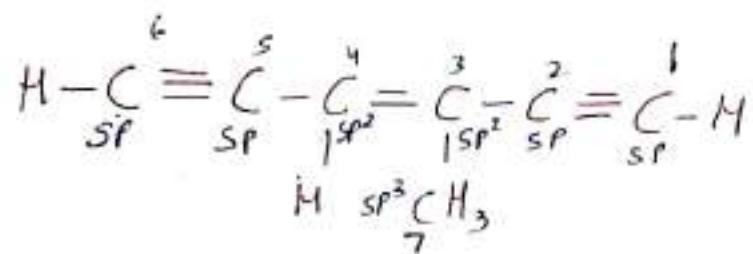
(47)



(48)



Q EN order?



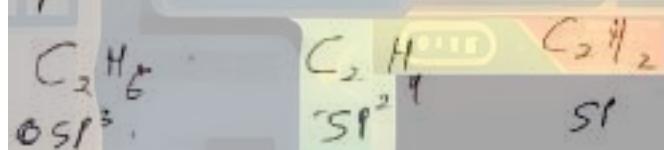
~~EN order of elements~~

EN q's character

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

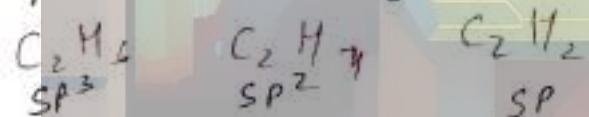
$(C_1 = C_2 = C_s = C_d) > (C_4 = C_3) > C_7$ [EN order]

Q 2. Compare C-C bond strength



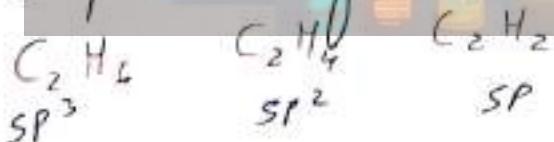
$C_2H_6 < C_2H_4 < C_2H_2$

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 mol. P on central atom).

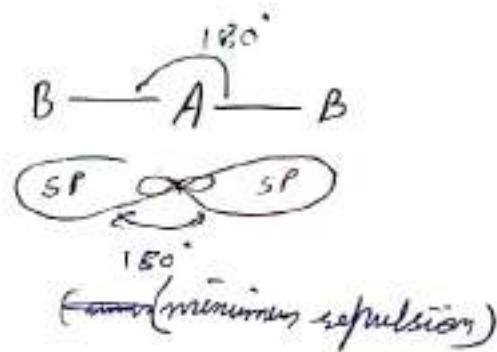
① SP-Hybridisation -

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

$S + (p_x/p_y/p_z)$

Geometry: linear (planar)

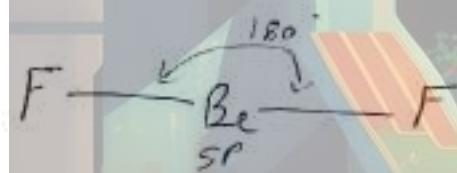
Angle: 180°



E.g. CO_2



E.g. ② BeF_2

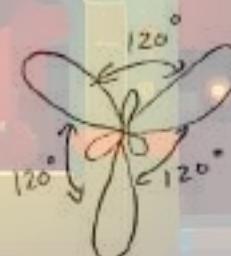


② SP^2 -Hybridisation -

$S + (\text{any } 2 p \text{ orbital})$

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

$S + p + P$

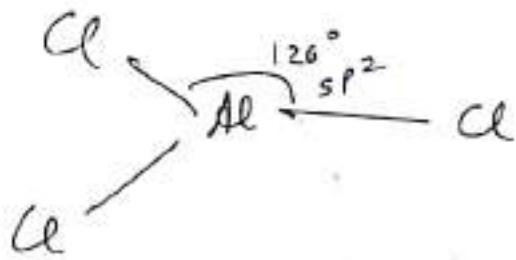


Geometry: Trigonal Planar

Bond angle: 120° ($^{\circ}$)

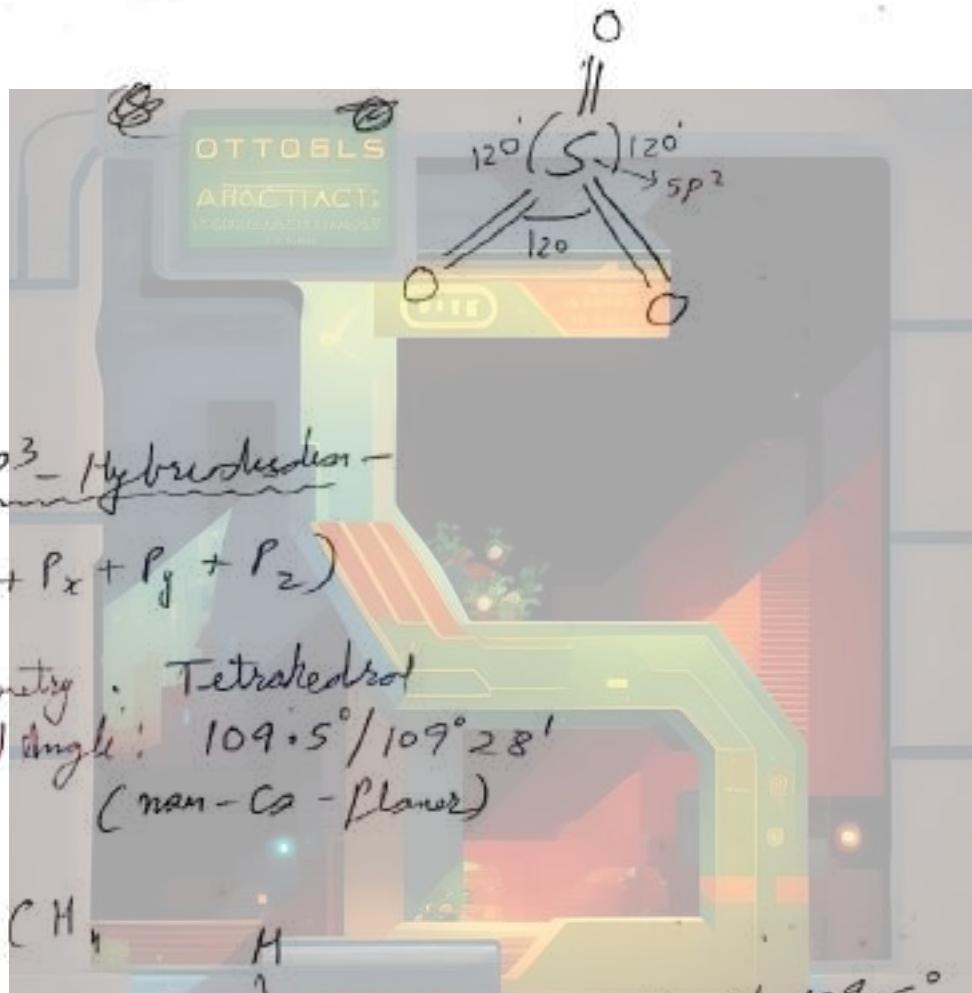
(Planar)

Eg ① AlCl_3



② SO_3

~~sp~~



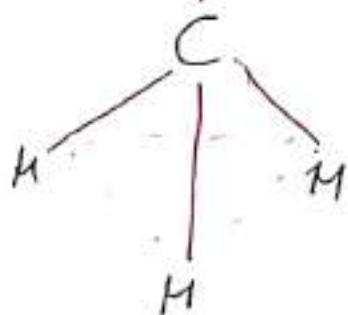
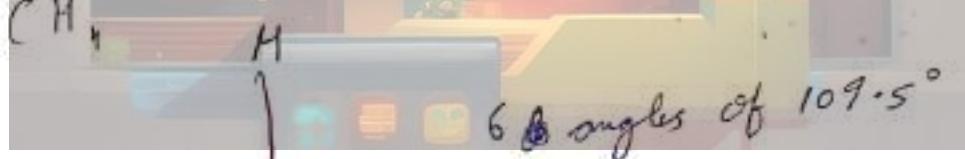
③ sp^3 -Hybridisation -

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

Geometry : Tetrahedral

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

Eg ①. CH_4



(52)

④ sp^3d Hybridization

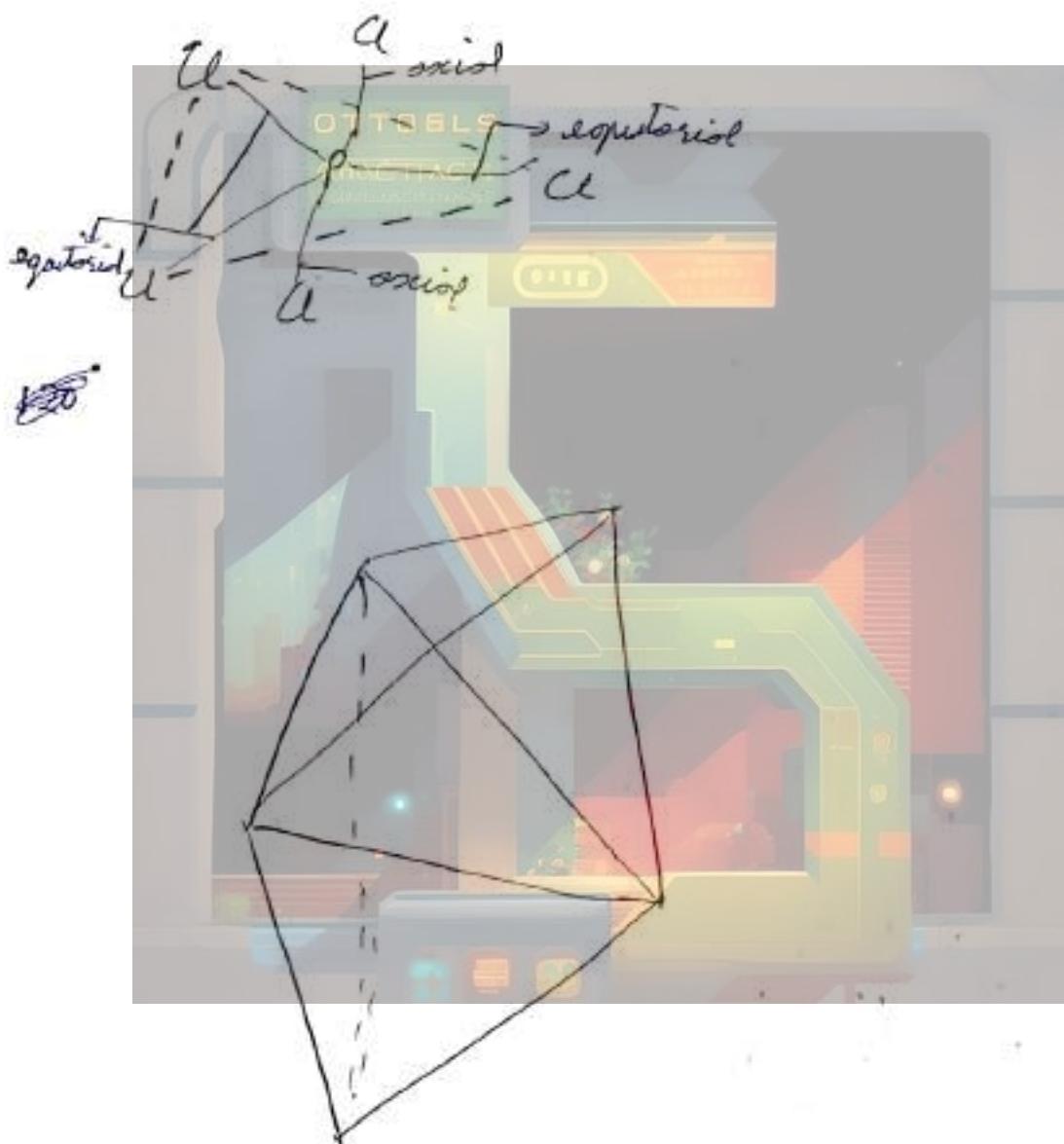
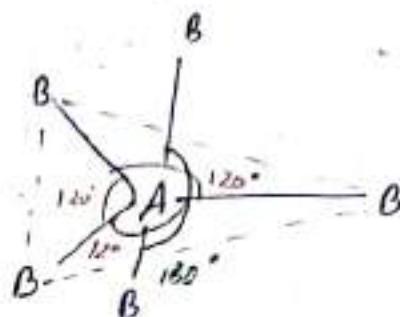
$$S + P_x + P_y + P_z + d_{z^2}$$

Geometry - Trigonal Bipyramidal

Bond Angl - $120^\circ(3)$
 $90^\circ(6)$

Non-Planar
 $180^\circ(1)$

Eg ① PCl_5



(5) SP^3d^2 Hybridisation

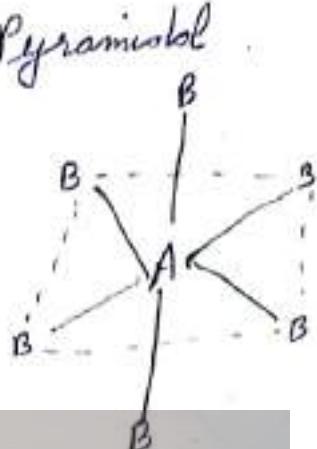
$$(S + P_x + P_y + P_z + d_{z^2} + d_{x^2-y^2})$$

Geometry :- Octahedral / Square Bi Pyramidal

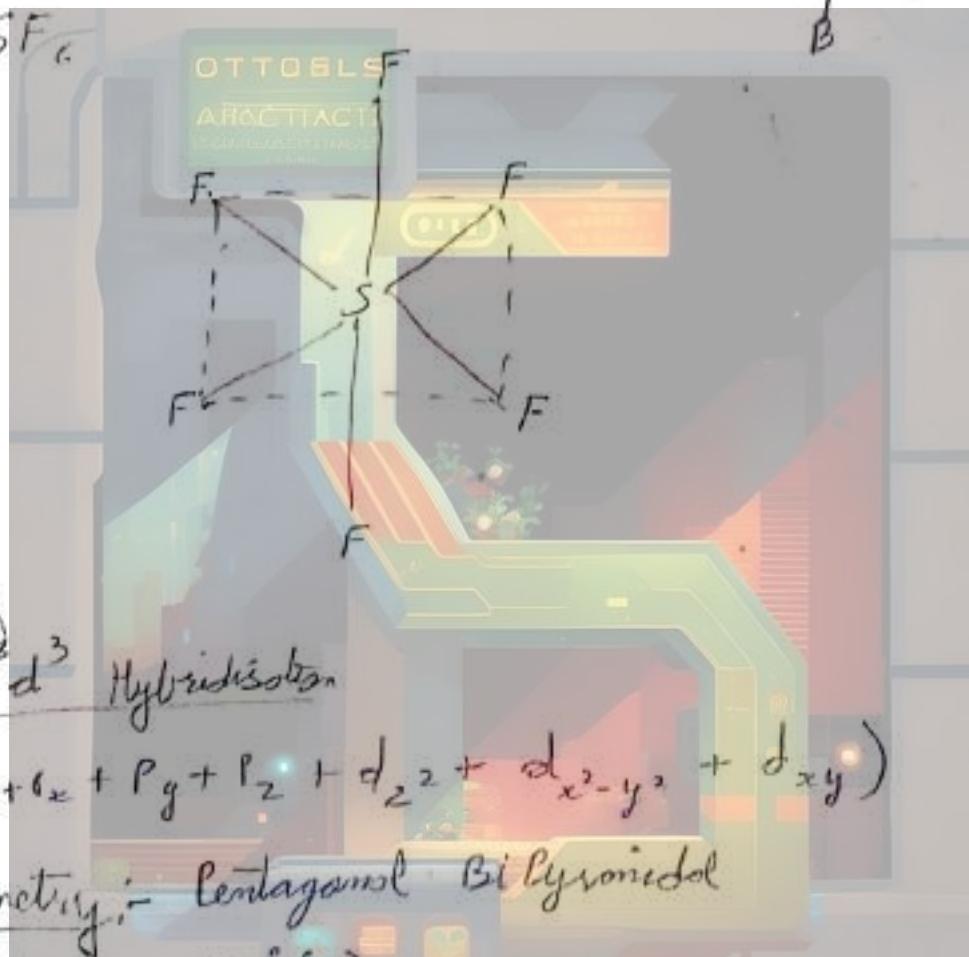
Band Angle :- $90^\circ(12)$

$180^\circ(3)$

Non-Planar



Eg 1. SF_6



(6) SP^3d^3 Hybridisation

$$(S + S_x + P_y + P_z + d_{z^2} + d_{x^2-y^2} + d_{xy})$$

Geometry :- Pentagonal Bi Pyramidal

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

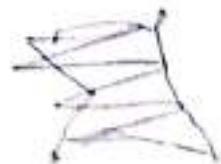
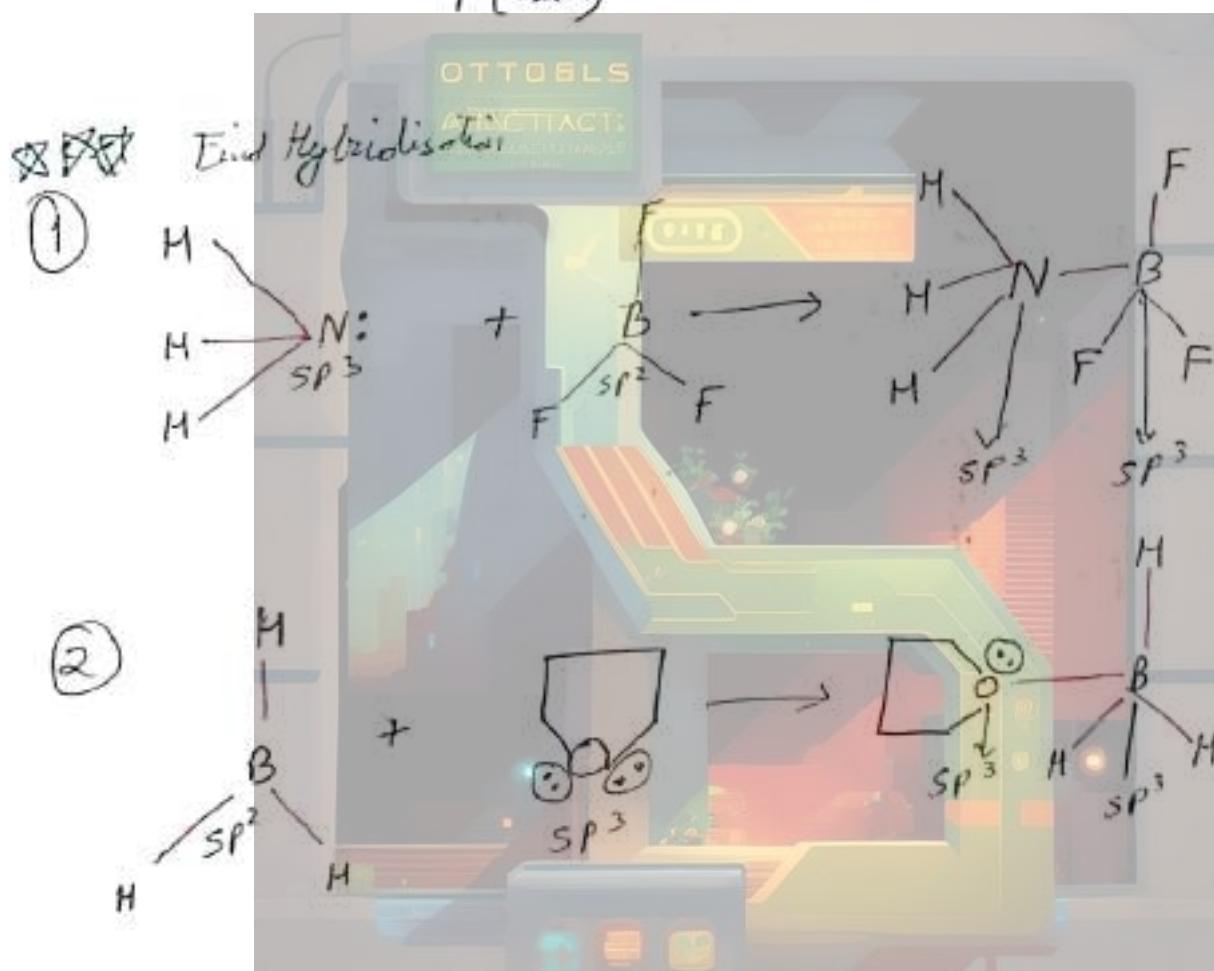
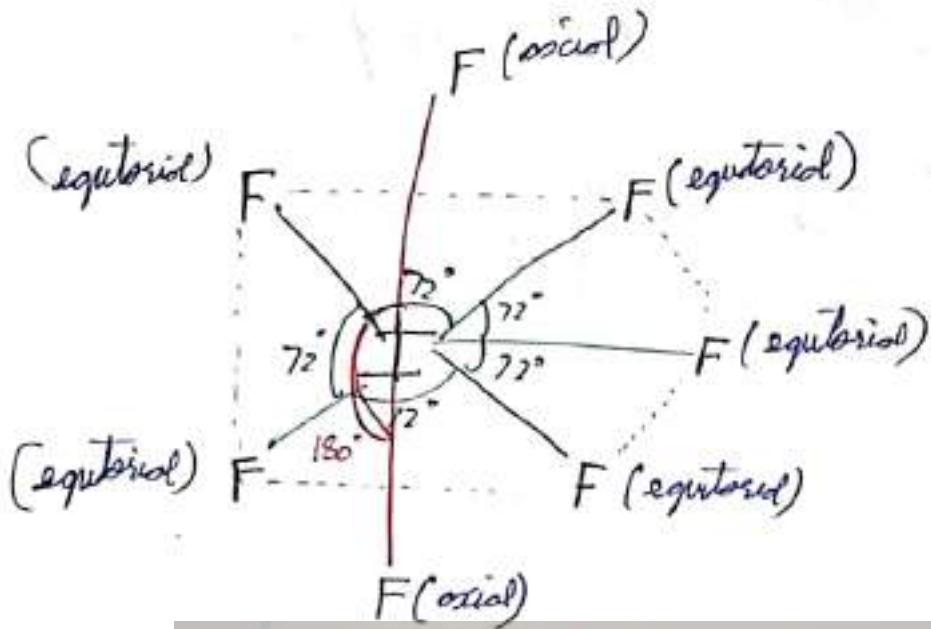
$90^\circ(10)$

$72^\circ(5)$

$19^\circ(5)$

Non-Planar

(54)



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

- a) d_{xy} b) d_{z^2} c) $d_{x^2-y^2}$ d) both b & c

Q Select the correct statement.

- ✓ 1) sp^2
✓ 2) sp^3

3) $n\text{s} + \text{any } 2\text{-nd} + 3\text{-rd} \rightarrow \text{sp}^3\text{d}^2$

4) $n\text{s} + 3\text{-rd} + 2(n-1)\text{d} \xrightarrow{\text{OTTOELS}} \text{sp}^3\text{d}^2$

5) $n\text{s} + 3\text{-rd} + 2 \text{ orbit}(n-1)\text{d} \rightarrow \text{sp}^3\text{d}^2$

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:

specie

- a) SO_4^{2-}
b) SO_3 (xz-None)
c) XeO_2F_2
d) XeF_6
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}

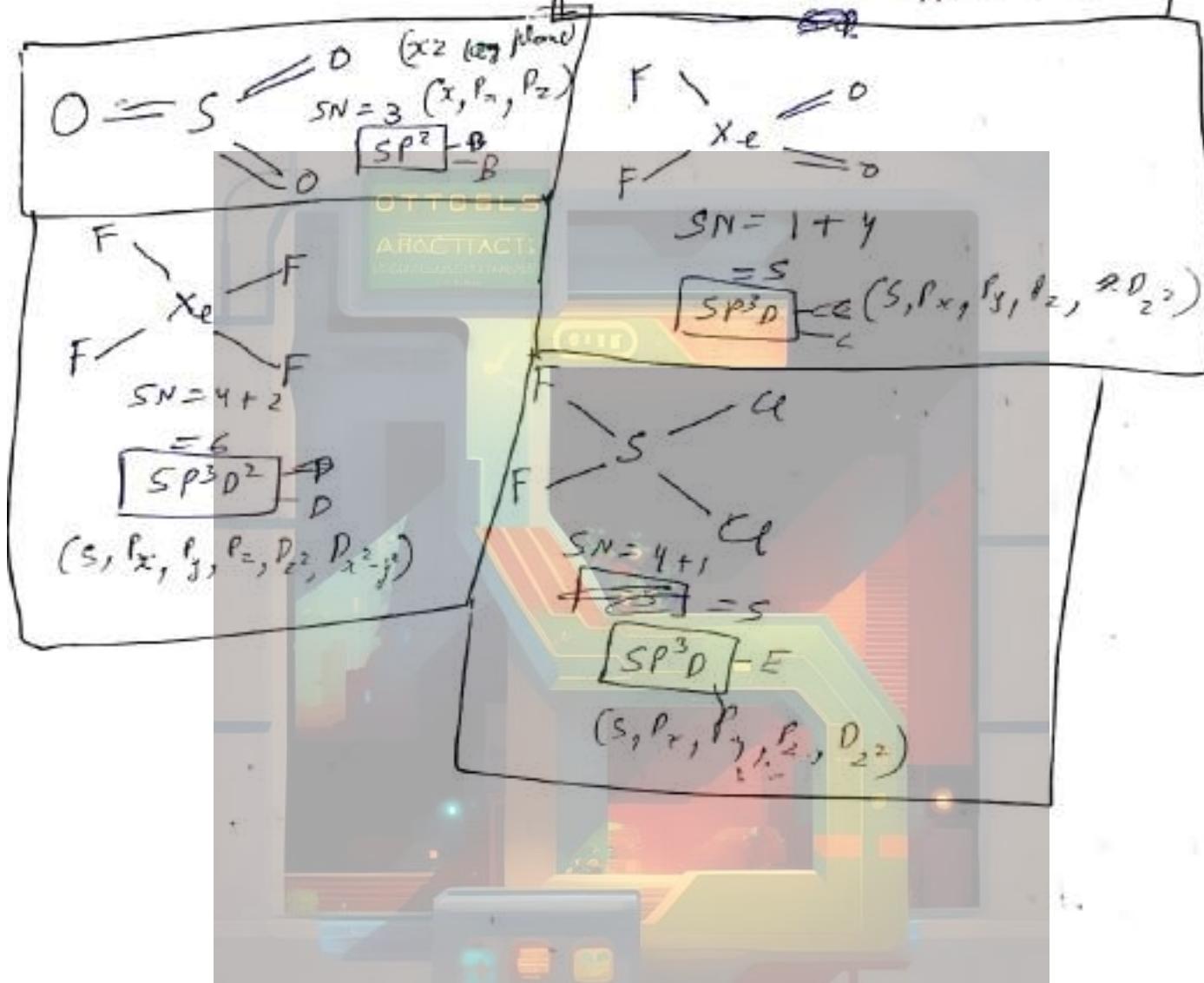
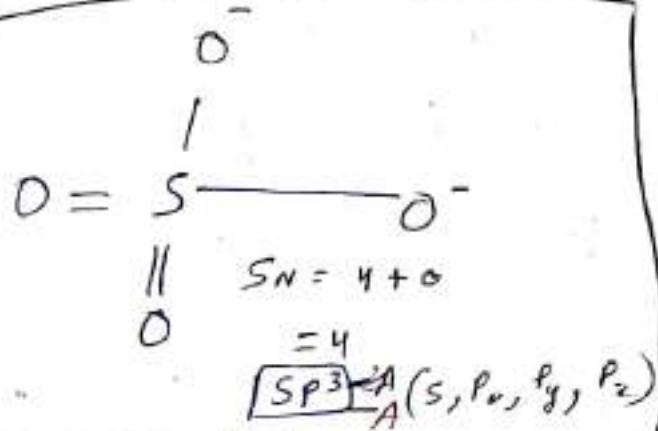
A - PQ

B - PQ

C - PQT

D - PQTS

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.

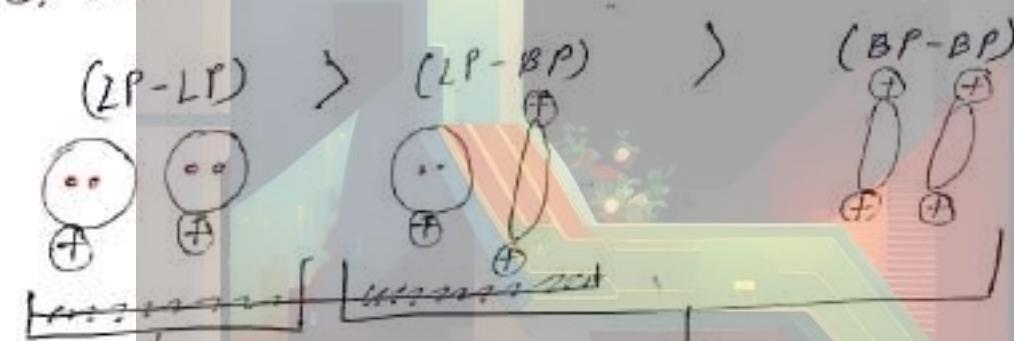
→ Minimum repulsion results in greater stability.

→ Repulsion is of 3 types.

1. LP-LP

2. LP-BP

3. BP-BP



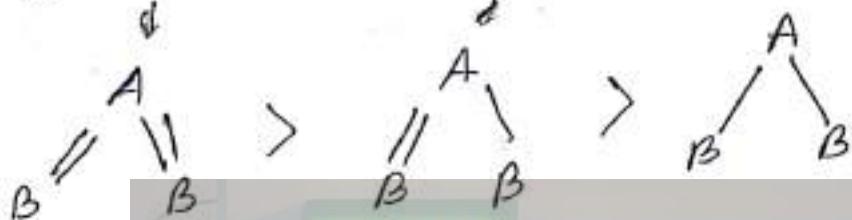
- Here the size of Lone Pair & e⁻ cloud is larger due to absence of 2nd nucleus.
→ Hence the distance b/w two lone pair is less & repulsion is high

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

Types of Bond Pair Repulsion

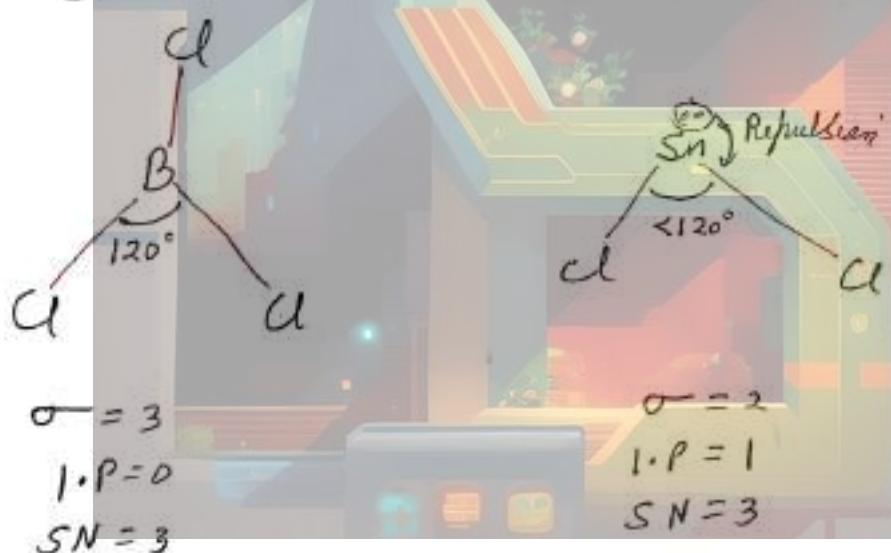
- ① Multiple Bond - Multiple Bond
- ② Multiple Bond - Single Bond
- ③ Single Bond - Single Bond

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

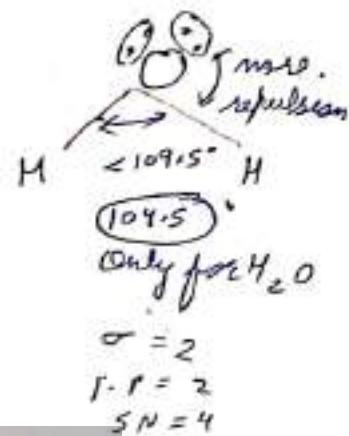
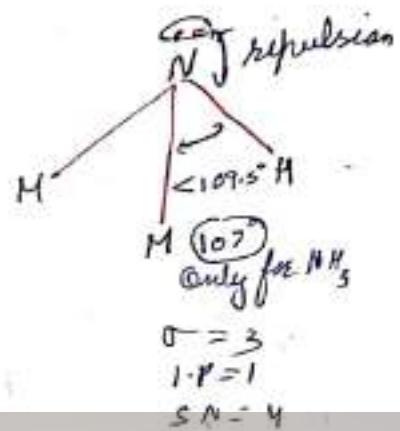
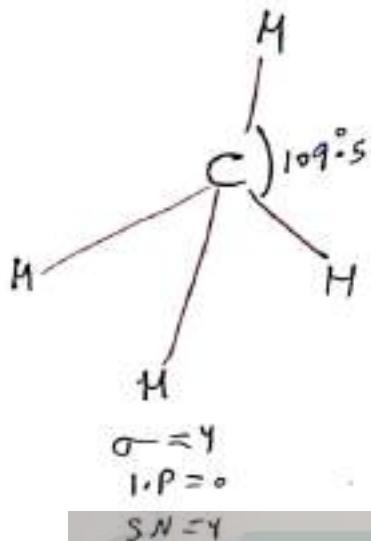


→ In Multiple bond, π electrons occupy more space around central atom so their repulsion is higher as compared to single bond.

Eg ① SP^2



② sp^3



→ 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 also decreases, similarly H_2O has 2 lone Pairs (L.P) & hence more repulsion so decrease in Bond Angle will be more.

③ sp^3d



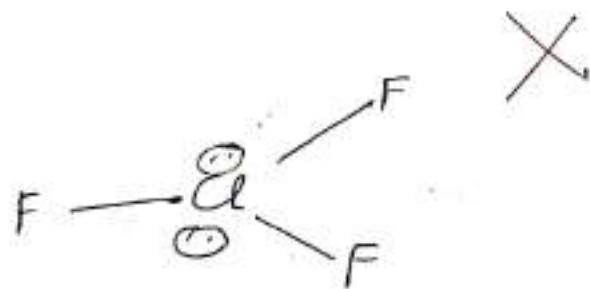
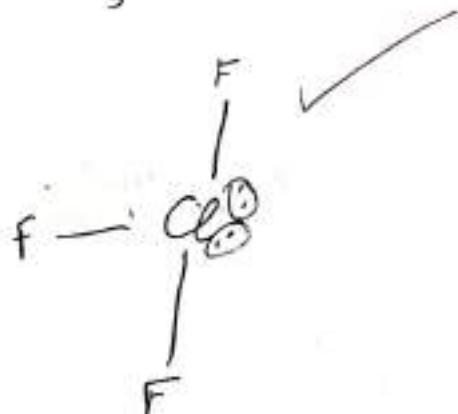
SF_4 ($\sigma = 4$, $L \cdot P = 1$)

at 90°
 $L \cdot P - B \cdot P = 2$

↓ less repulsion

at 90°
 $L \cdot P - B \cdot P = 3$

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



at 90°

$$\text{LP} - \text{BP} = 6$$

at 90°

$$\text{LP} - \text{BP} = 4$$

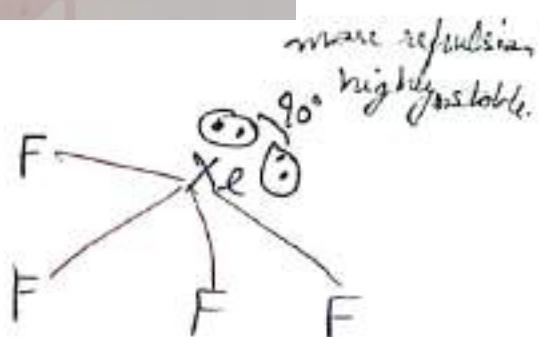
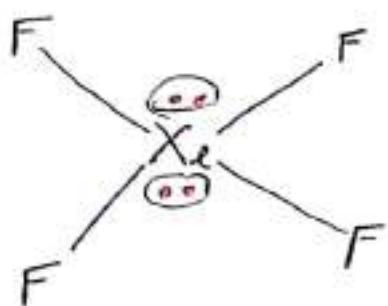
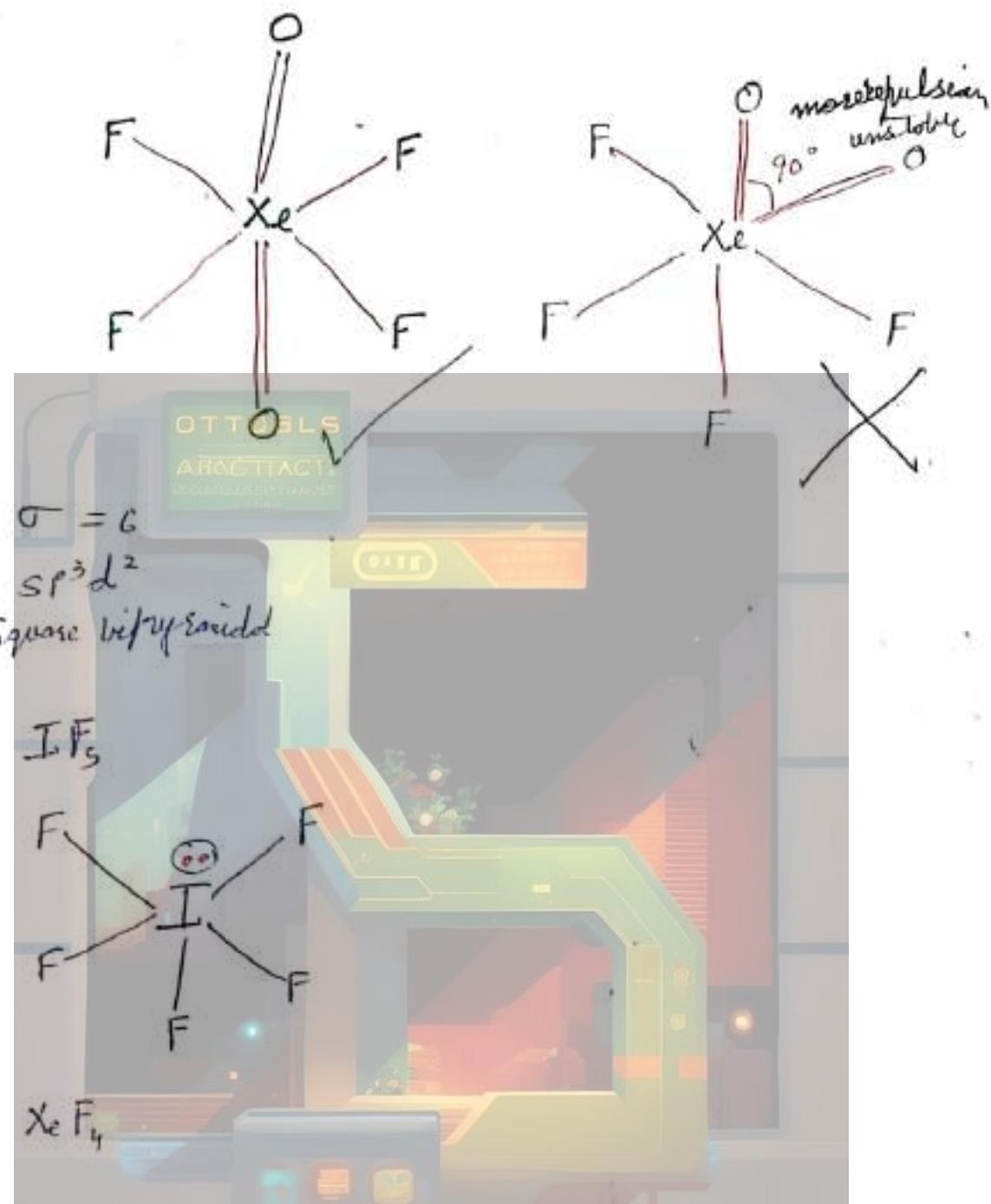
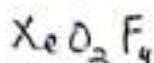
less Repulsion

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

~~XeO~~ XeO_3F_2



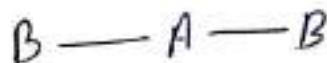
(4) SF_6 (Put lone pair & double bond at 180°)



A 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^2



① σ-bond = 2

② Lone pair = 0

③ No. of H.O. = 2

④ Hybridisation = SP^2

⑤ e⁻-geometry = linear

⑥ Shape = linear

⑦ Bond angle = 180° (±)

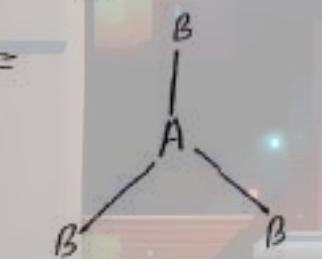
⑧ Planar

⑨ & max no. of atoms in one plane = 3

SP^2



~~AB₃~~



① σ = 3

② L.P. = 0

③ H.O. = 3

④ SP^2

⑤ Trigonal Planar

⑥ Trigonal Planar

⑦ Bond angle = 120° (3)

⑧ Planar

⑨ 4

(63)

③ AB_3E , AB_3 represent a lone pair
e.g. $SOCl_2$



① $\sigma = 2$

② $L.P = 1$

③ $H.O = 3$

④ SP^2

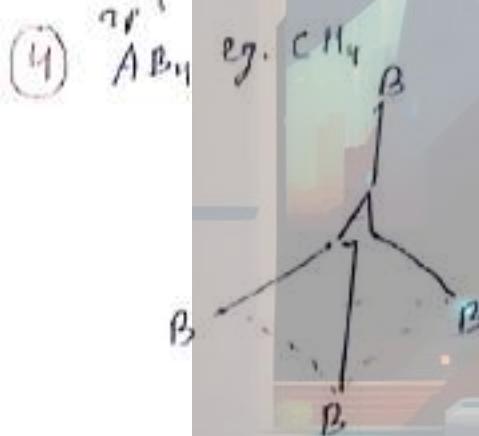
⑤ Trigonal Planar

⑥ V-shape / Bent Shape / Angular

⑦ $<120^\circ (1)$

⑧ Planar

⑨ 3



① $\sigma = 4$

② $L.P = 0$

③ $H.O = 4$

④ SP^3

⑤ Tetrahedral

⑥ Tetrahedral

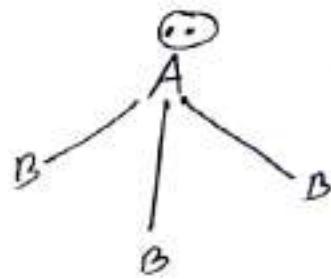
⑦ $109.5^\circ (4)$

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

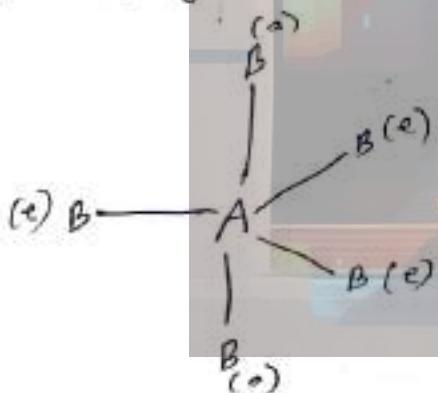
⑦ $109^\circ 28'$ $< 109.5^\circ (3)$

⑧ Non-Planar

⑨ $B = 3$

SP^3D

⑦ AB_5 eg. PCl_5



① $\sigma = 5$

② $L.P = 0$

③ $H.O = 5$

④ SP^3D

⑤ Trigonal Bifurconical

⑥ Trigonal Bifurconical

⑦ $120^\circ (3)$

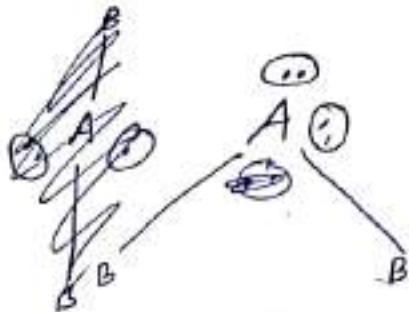
$90^\circ (e)$

$180^\circ (1)$

⑧ Non-Planar

⑨ $B = 4$

⑥ AB_2E_2 eg. H_2O



① $\sigma = 2$

② $L.P = 2$

③ $H.O = 4$

④ SP^3

⑤ Tetrahedral

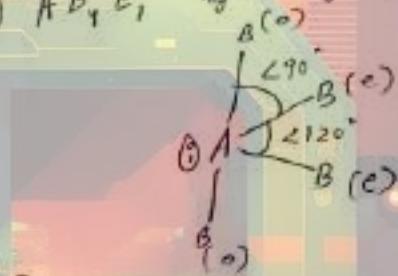
⑥ ~~Tetrahedral~~ V-Shape

⑦ $< 109.5^\circ (1)$

⑧ Non-Planar

⑨ $B = 3$

③ AB_3E_1 eg. SF_4



① $\sigma = 4$

② $L.P = 1$

③ $H.O = 5$

④ SP^3D

⑤ Trigonal Bifurconical

⑥ ~~Tetrahedral~~ See-Saw

⑦ $120^\circ (4)$

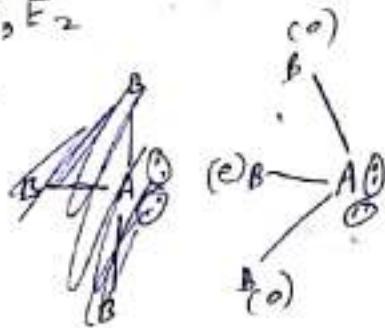
~~$< 109.5^\circ (1)$~~ $< 180^\circ (1)$

⑧ Non-Planar

⑨ $B = 3$

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

⑥ Distorted T-Shaped

⑦ $< 90^\circ (2)$

$\leq 180^\circ (1) \text{ a-a}$
FACT

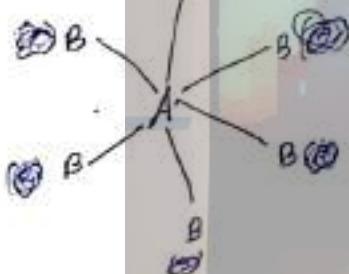
⑧ Non-Planar

⑨ 4

SP^3D^2

⑩ AB_5

e.g. SF_6



① $\sigma = 6$

② $L \cdot P = 0$

③ $H \cdot O = 6$

④ SP^3D^2

⑤ Square BiPyramidal

⑥ Square BiPyramidal

⑦ $90^\circ (12)$

$180^\circ (2)$

⑧ Non-Planar

⑨ 5

⑩ $AB_2 E_3$ e.g. XeF_2



① $\sigma = 2$

② $L \cdot P = 3$

③ $H \cdot O = 5$

④ SP^3D

⑤ Trigonal BiPyramidal

⑥ Linear

⑦ $\geq 180^\circ (1) \text{ a-a}$

⑧ Planar

⑨ 3

⑩ $AB_5 E_1$

e.g. $Br_2 F_5$

(a) B

(c) B

(b) B

(a) B

(c) B

(b) B

① $\sigma = 5$

② $L \cdot P = 1$

③ $H \cdot O = 6$

④ SP^3D^2

⑤ Square BiPyramidal

⑥ Square Pyramidal

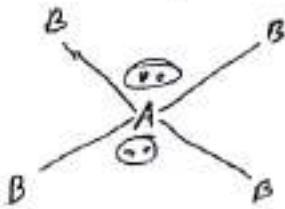
⑦ $< 90^\circ (8) - \text{around } 70^\circ$

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

⑧ Non-Planar

⑨ 5

(13) AB_3E_2 eg XeF_4



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

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

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

$\textcircled{5}$ Square Pyramidal

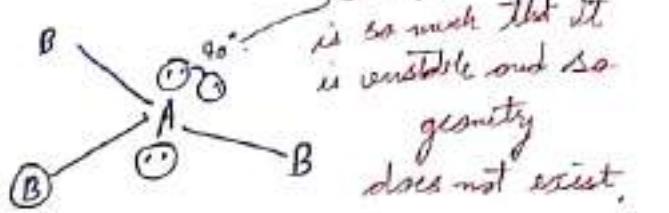
$\textcircled{6}$ Square Planar

$\textcircled{7}$ $90^\circ(4)$ $180^\circ(2)$

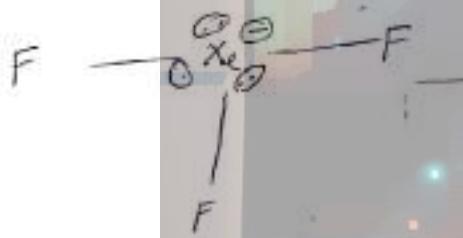
$\textcircled{8}$ Planar

$\textcircled{9}$ 5

(14) AB_3E_3



$\textcircled{10}$ XeF_3^- does not exist why



$$L.P = 3$$

$$\sigma = 3$$

$$H.O = 6$$

$$SP^3D^2$$

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

90° - very high repulsion.

(67)

H.W. 04-06-2024

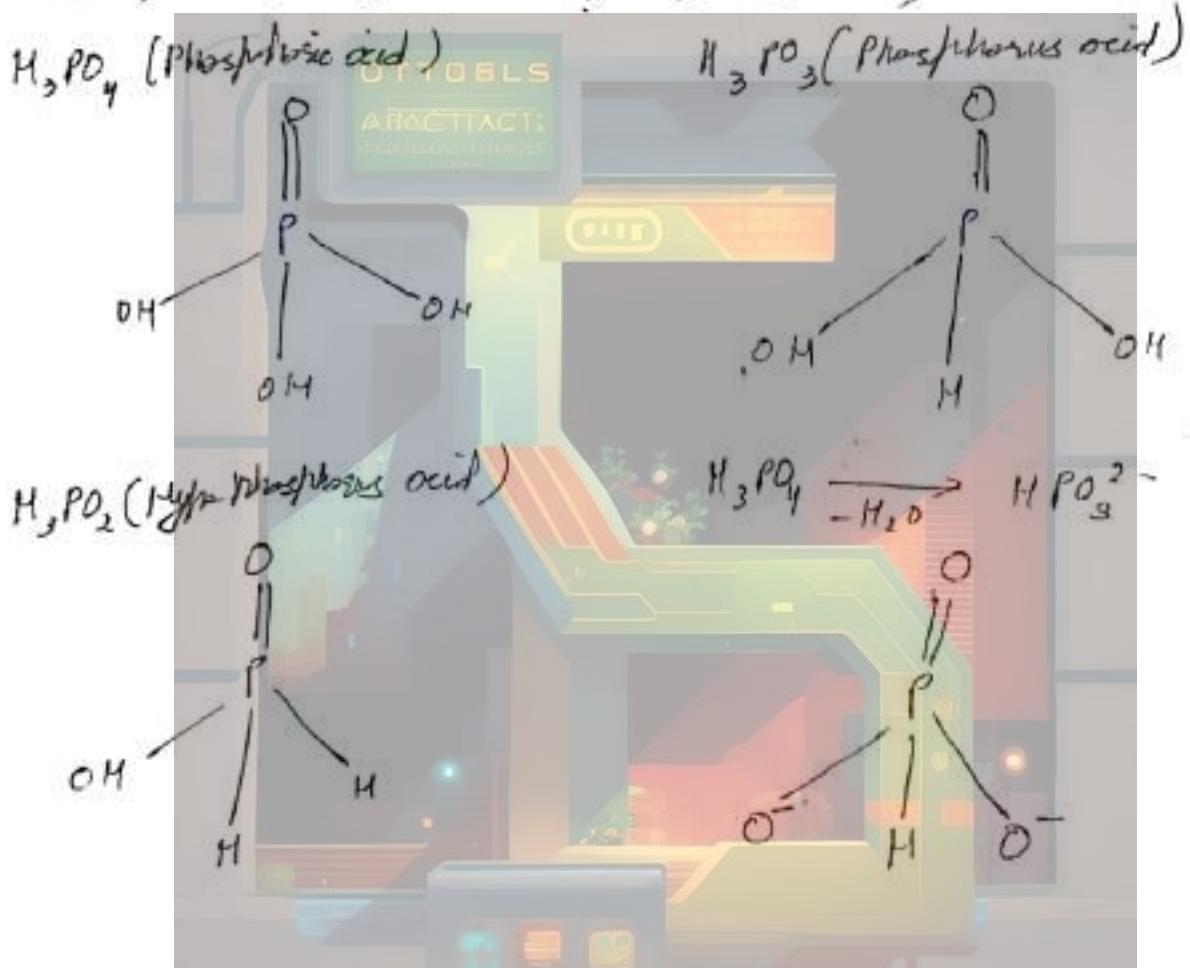
D-1 (8-11)

D-2 (6-11)

Roc - 13, 14

H.W. 06-06-2024

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

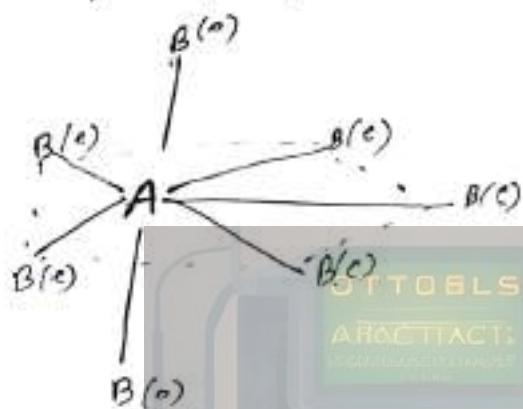


~~H~~

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

SP^3D^3

(15) AB_7 e.g. IF_7



$$\textcircled{1} \sigma = 7$$

$$\textcircled{2} L.P = 0$$

$$\textcircled{3} H.O = 7$$

$$\textcircled{4} s SP^3D^3$$

\textcircled{5} Pentagonal Bi Pyramidal

\textcircled{6} Pentagonal Bi Bi Pyramidal

$$\textcircled{7} 72^\circ(5)-(e-e)$$

$$90^\circ(10)-(o-o)$$

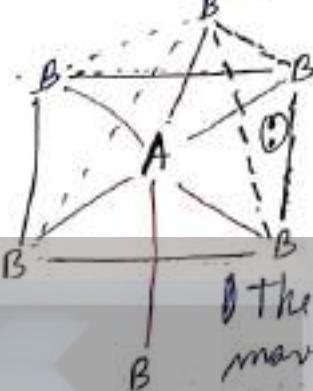
$$180^\circ(1)-(e-e)$$

\textcircled{8} Non-Planar

\textcircled{9} 6

~~Xood Raksha Hai~~

(16) AAB_6E e.g. $XeF_6, IF_5, XeOF_5^-$



The lone pair keeps moving in the block
Distorted Octahedral

$$\textcircled{1} \sigma = 6$$

$$\textcircled{2} L.P = 1$$

$$\textcircled{3} H.O = 7$$

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

\textcircled{5} Pentagonal Bi Pyramidal

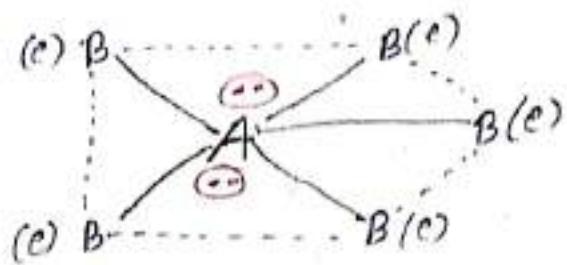
\textcircled{6} Distorted Octahedral
or
Capped Octahedron

\textcircled{7} X

\textcircled{8} Non-Planar

\textcircled{9} X

(17) AB_5E_2 e.g. XeF_5



① $d = 5$

② $L \cdot P = 2$

③ $H \cdot O = 7$

④ SP^3D^3F

⑤ Bent and Bi-pyramidal

⑥ ~~Bent~~ Bent and Planar

SP^3D^3F $L \cdot P \rightarrow 2\pi$

180° \rightarrow 2π

⑦ $72^\circ(s) - (c-c)$
 $119^\circ(ss) - (c-c)$

⑧ Planar

⑨ 6

⑩ Draw Structure

① PCl_4^+

② I_3^-

③ SO_2

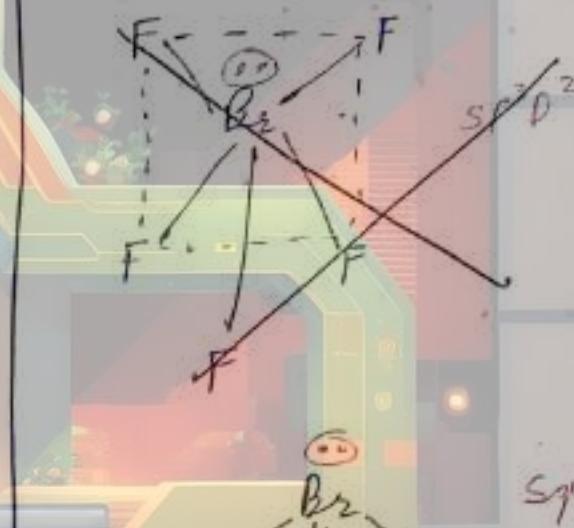
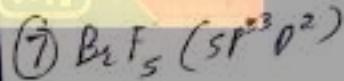
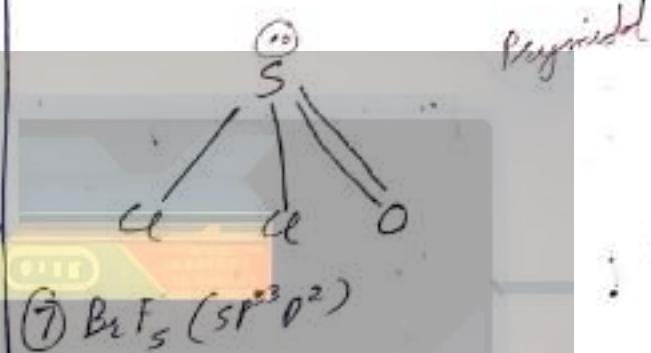
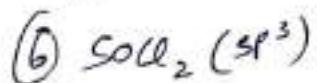
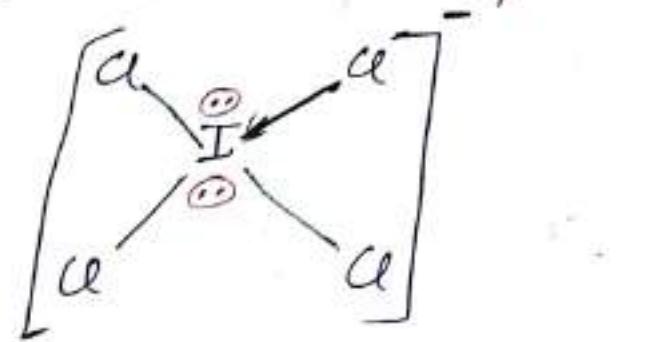
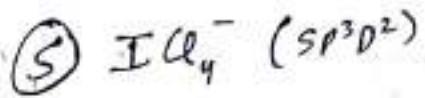
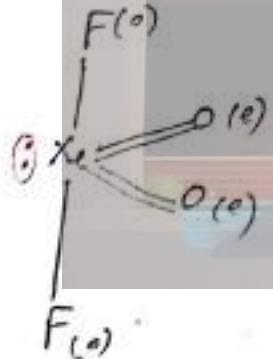
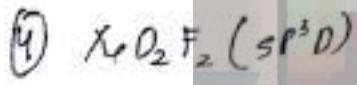
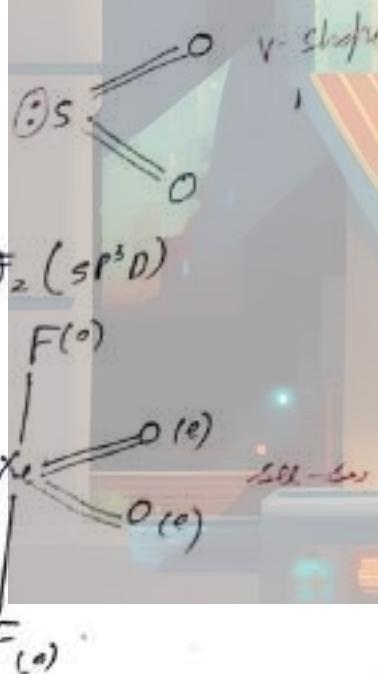
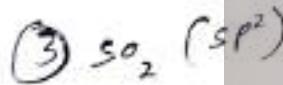
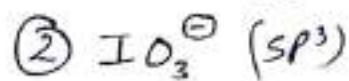
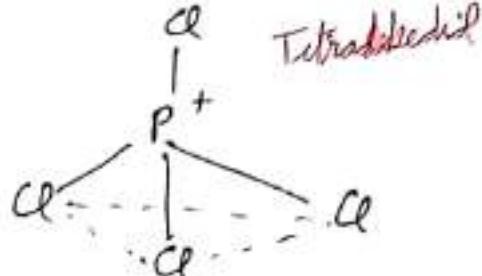
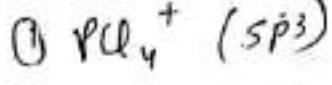
④ XeO_2F_2

⑤ ICl_4^-

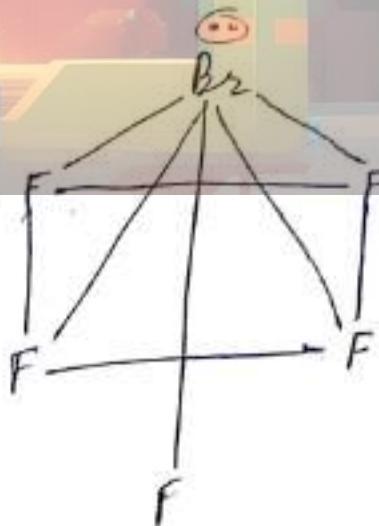
⑥ $SOCl_2$

⑦ B_2F_5

⑧

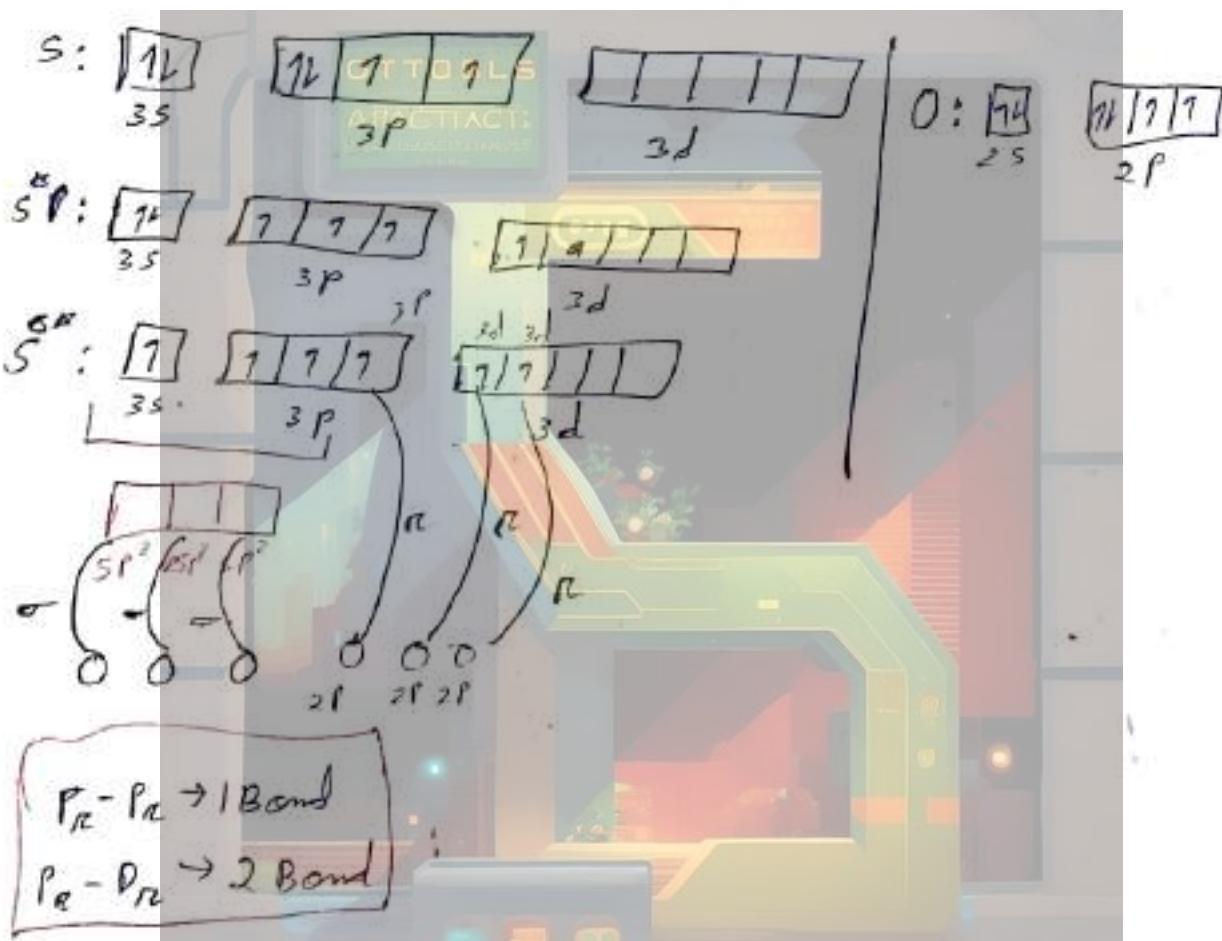
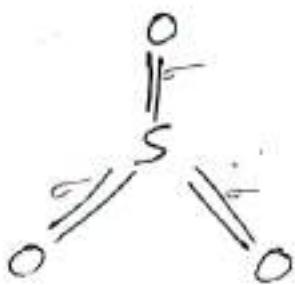


Square Pyramidal



Q. SO_3 How many π -bonded types of π -bonds
Type of π -bond calculation

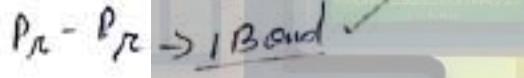
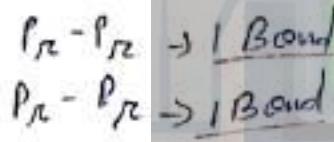
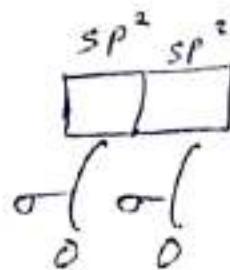
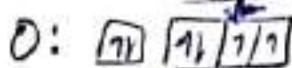
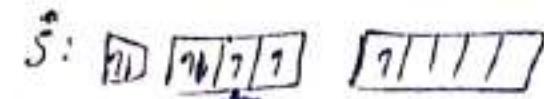
SP^2



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

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

① SO_2



② ClO_2^{\oplus}

③ XeO_4

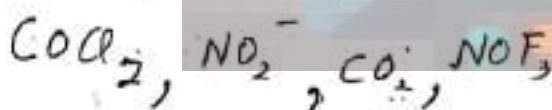
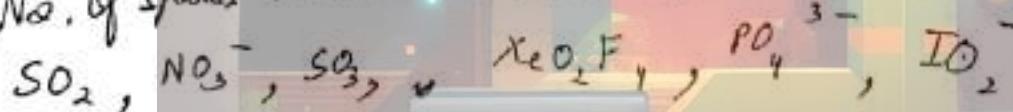
④ ClO_4^-

⑤ C_2H_2

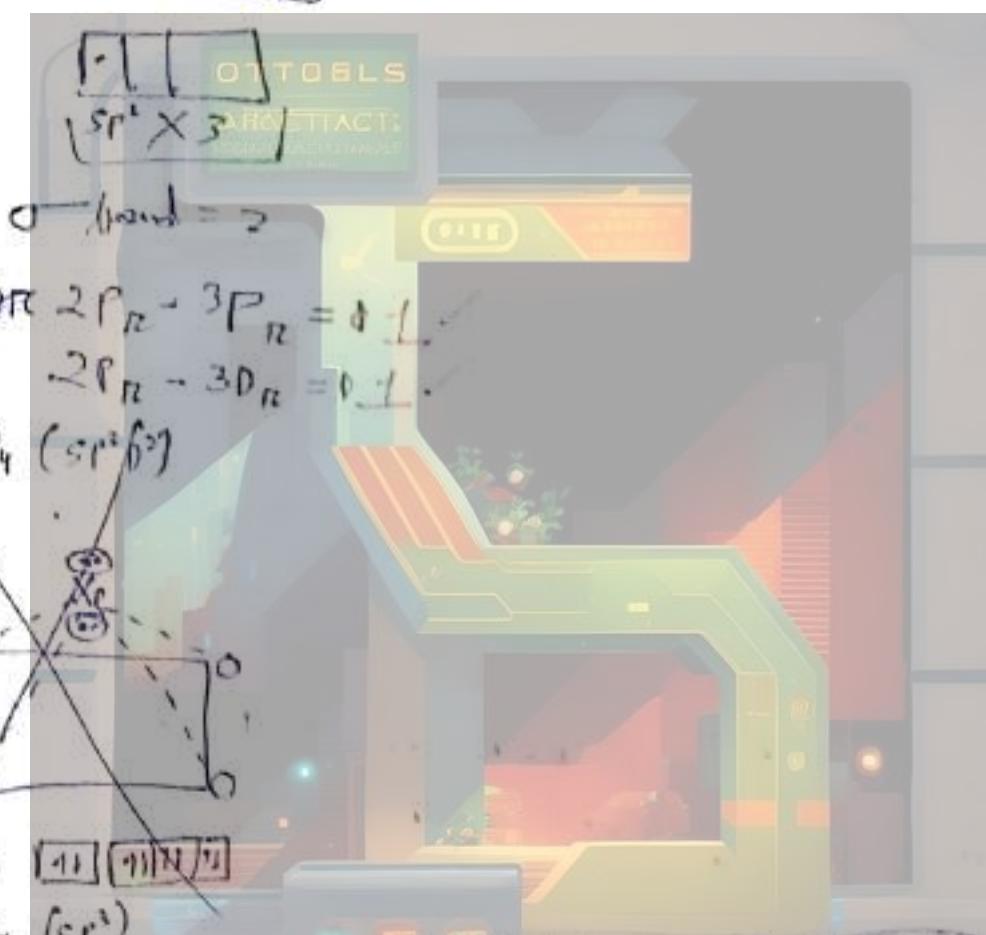
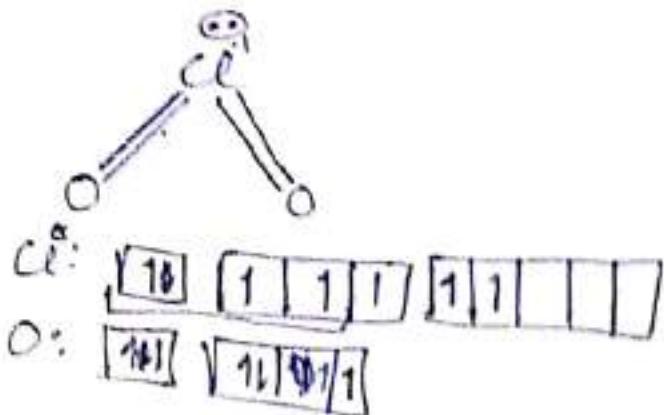
⑥ CO_3^{2-}

⑦ CO_3^-

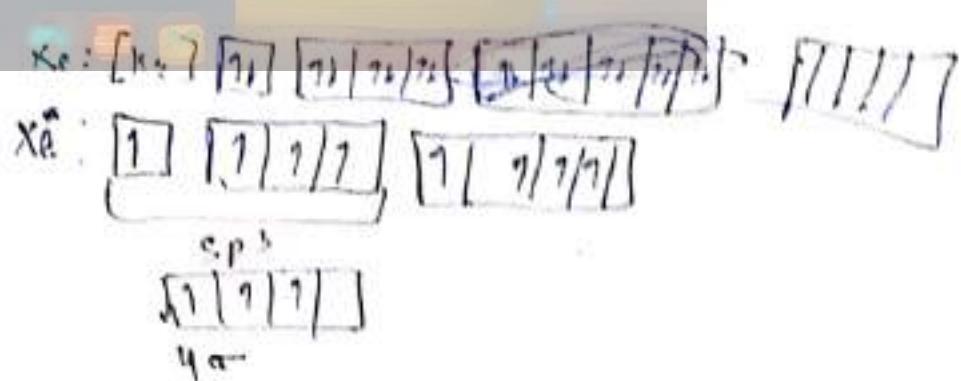
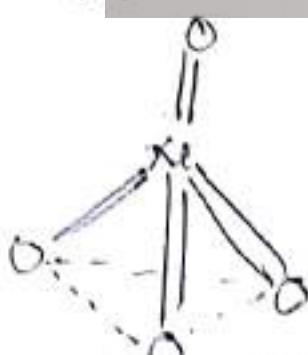
Q2. No. of species that have more than $\text{P}_{\text{R}} - \text{P}_{\text{R}}$ than $\text{P}_{\text{R}} - \text{P}_{\text{R}}$ bonds :-



Q1 ② $\text{ClO}_4^- (\text{sr}^*)$

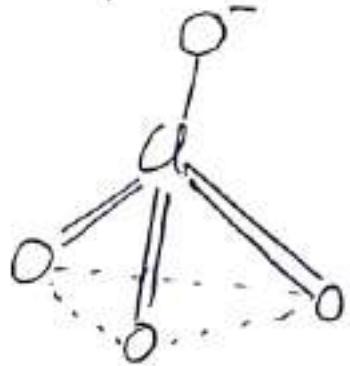
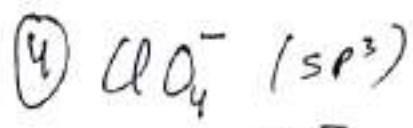


(3) $\text{XeO}_4 (\text{sr}^*)$

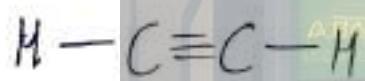
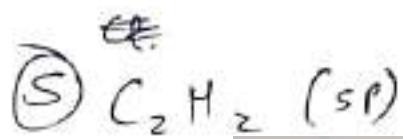


$$(2P_N - 5P_N) \quad P_N - P_N = 0 \downarrow$$

$$P_N - P_N = 0$$

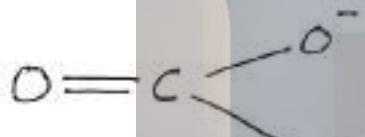
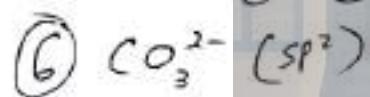


$$P_R - P_{R_d} = 0 \checkmark$$
$$P_R - P_{R_e} = 3 \checkmark$$



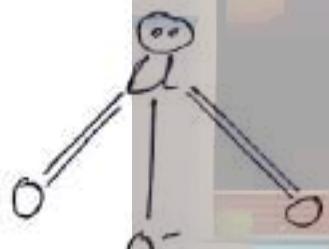
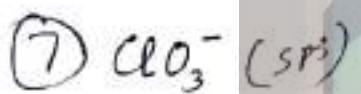
$$P_R - P_{R_d} = 2 \checkmark$$

$$P_R - P_{R_e} = 0 \checkmark$$



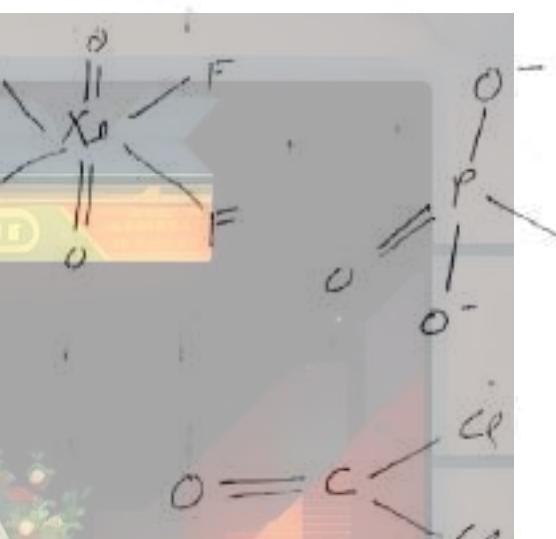
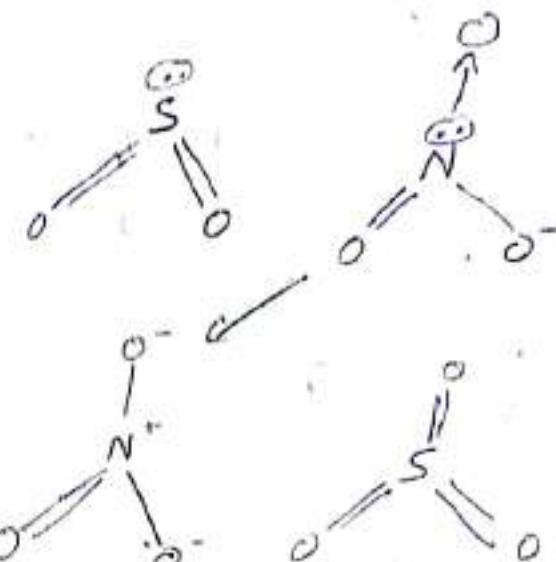
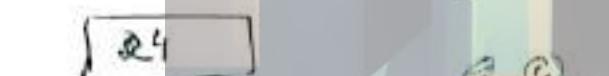
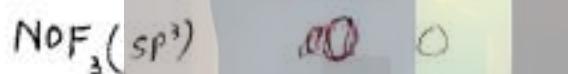
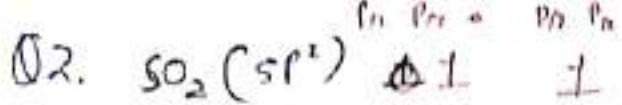
$$P_R - P_{R_d} = 1 \checkmark$$

$$P_R - P_{R_e} = 0 \checkmark$$

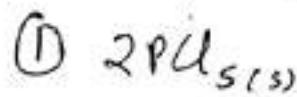


$$P_R - P_{R_d} = 0 \checkmark$$

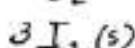
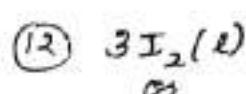
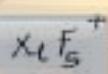
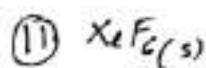
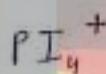
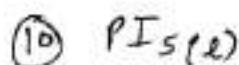
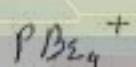
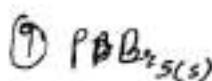
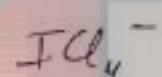
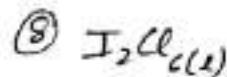
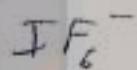
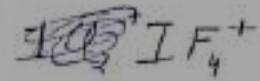
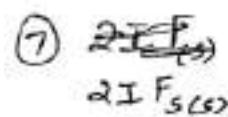
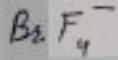
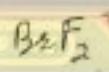
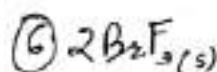
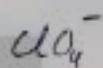
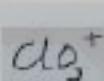
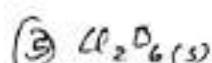
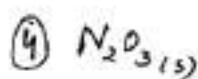
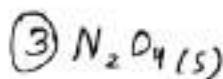
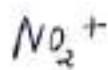
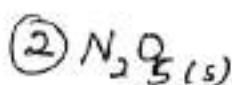
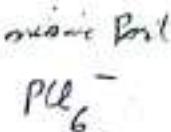
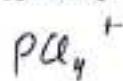
$$P_R - P_{R_e} = 2 \checkmark$$



Solid/Liquid state hybridisation (Rotha Mai)

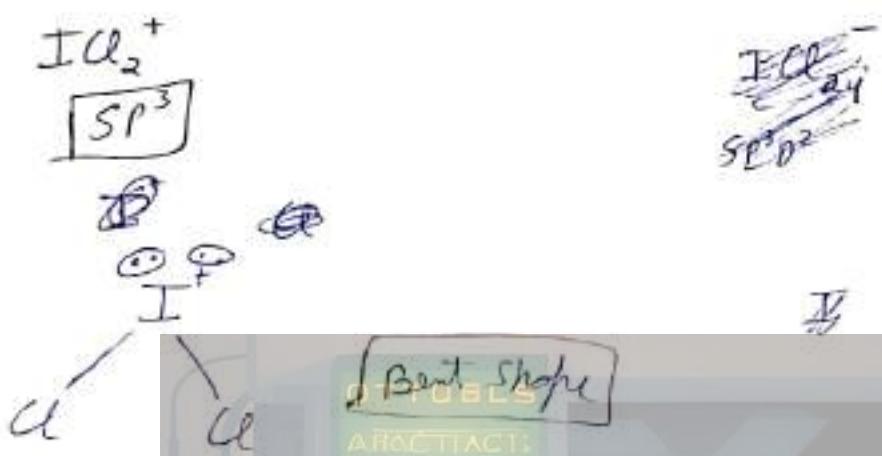


ionic Part



Q1

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



Bent shape

ABO₃T₃

Q2. What is the bond angle diff between anionic part of solid I_2Cl_6 & anionic part of N_2O_5

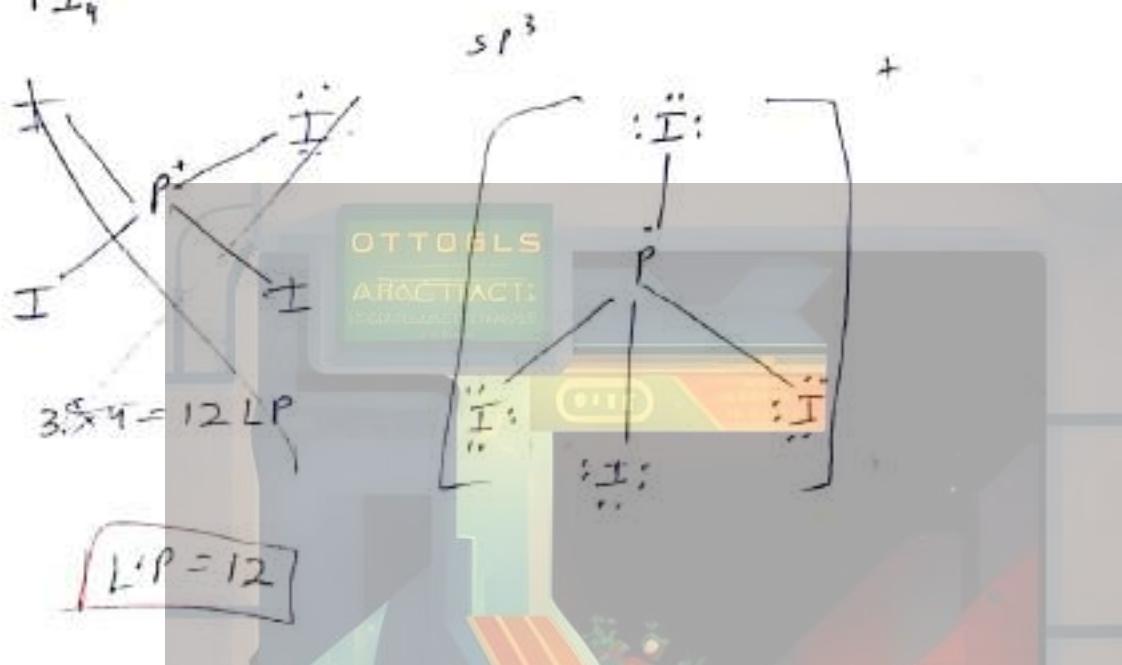
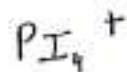


Q Comment hybridisation of anionic Part of $PBr_5(s)$



No hybridisation

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

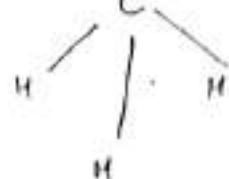


Bond Parameter :-

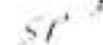
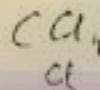
① Bond angle :-

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 in polarisability molecules)

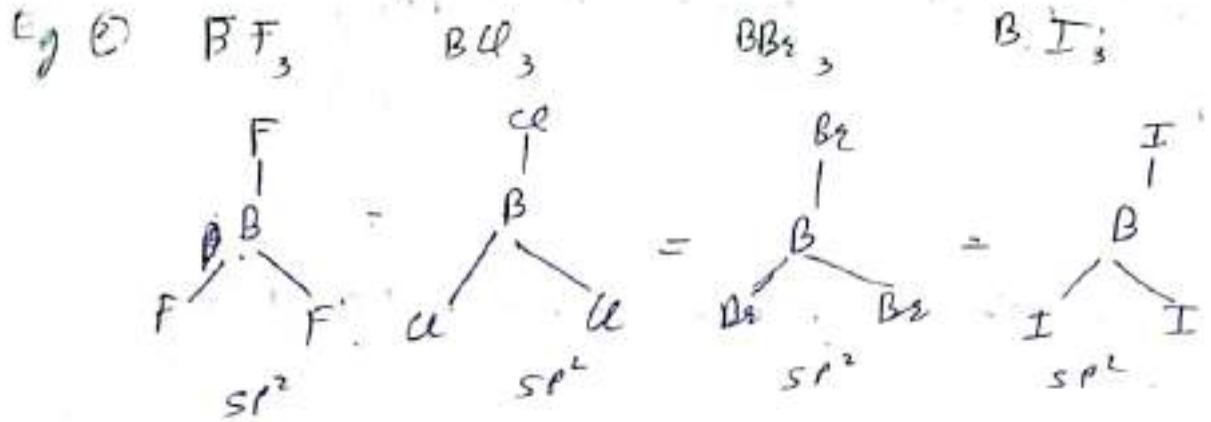
e.g.



$$L.P = 0$$



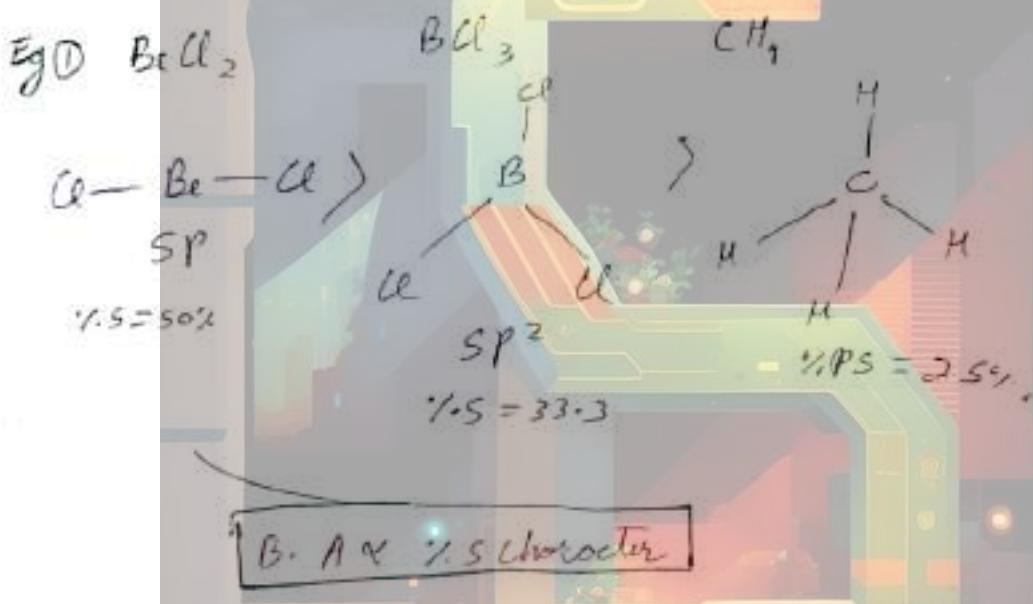
$$BA = \text{same} = 109.5^\circ$$



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

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

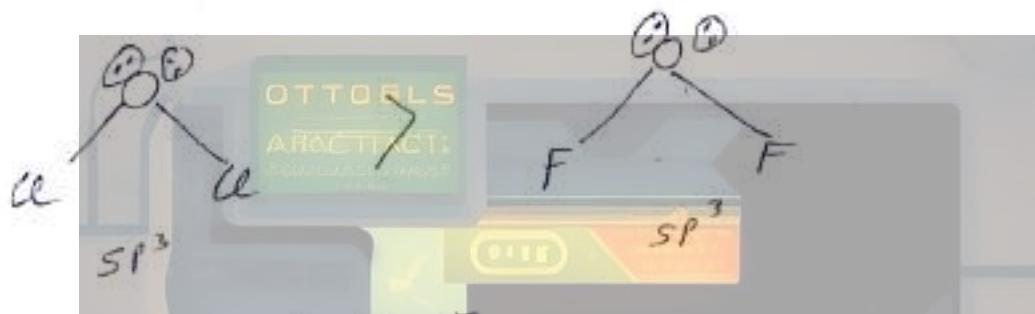
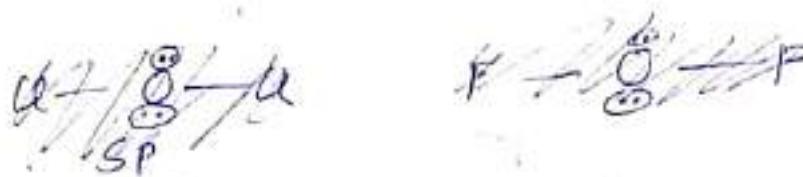
Rule 2 :- If hybridisation of central atom is different then the bond angle will be directly proportional to %s character.



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

Size of surrounding atom \uparrow B.A↑

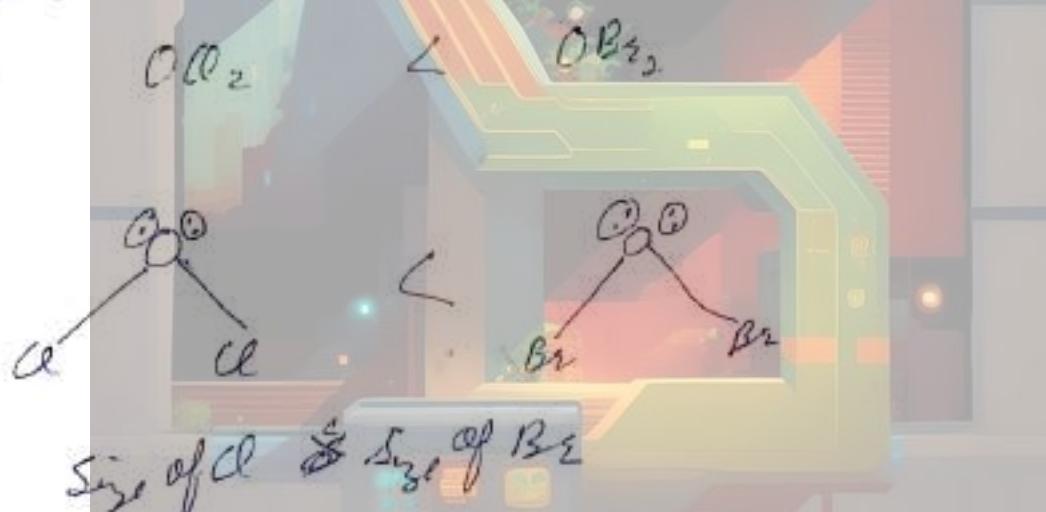
Note - If central atom has at least one lone pair and surrounding atoms belongs to 3rd period elements onwards like Cl, Br or halogen group (Cl¹⁻, C₆H₅Cl) Then steric crowding will appear.



Size of Cl > Size of F

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

Eg ②



Size of Cl < Size of Br

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

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 CH_4

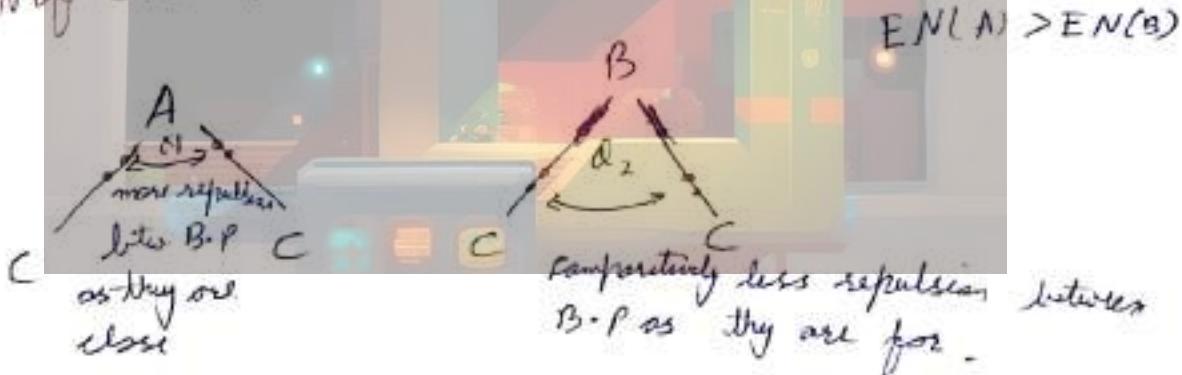
NH_3

H_2O



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

i) EN of Central Atm. (All the surrounding atoms are same)



Eg ① OCl_2



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

ii) EN of surrounding atom {Same ~~as~~ central atom?}



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

Eg

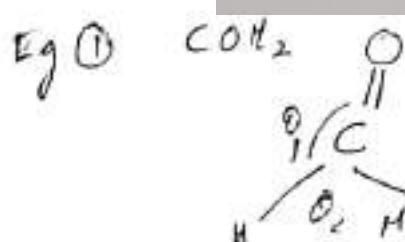


$\text{EN}(\text{N}) < \text{EN}(\text{O})$

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

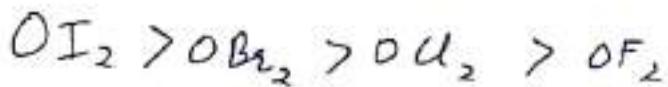
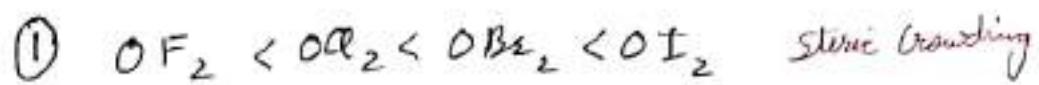
Rules - Multiple Bond repel more strongly as compared to single bond.

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



$\boxed{\theta_1 > \theta_2}$

Q Bond angle order?



②

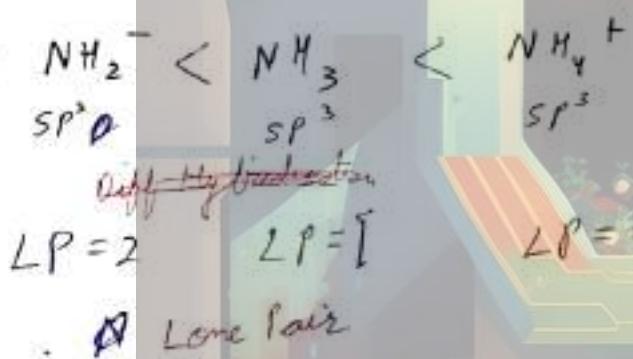


O

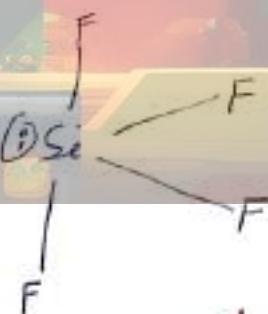
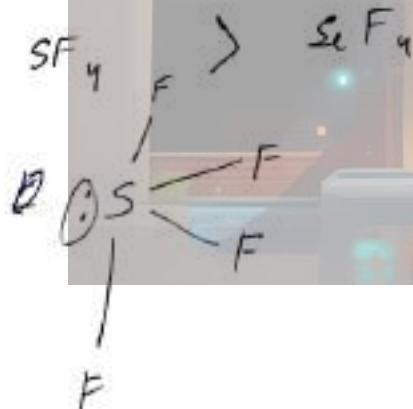
O

EN of surrounding atom

③

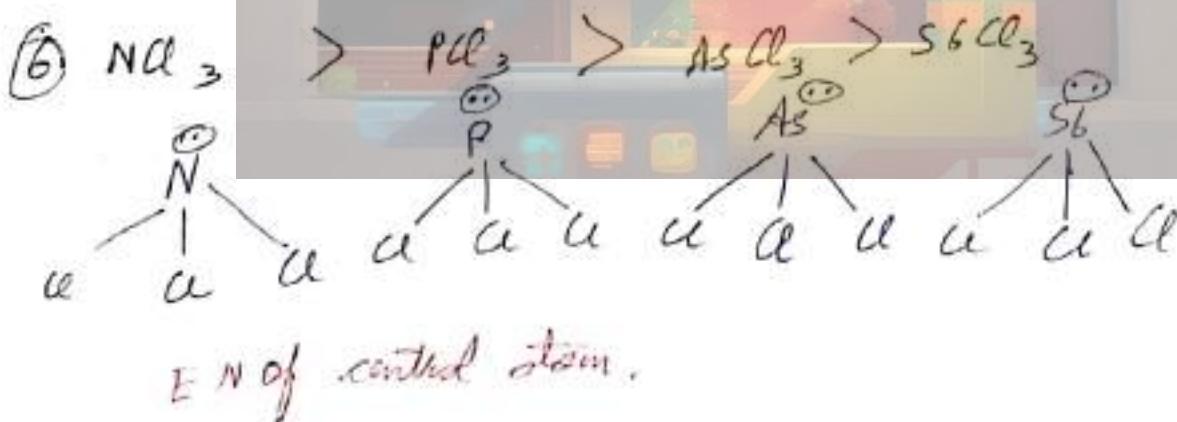
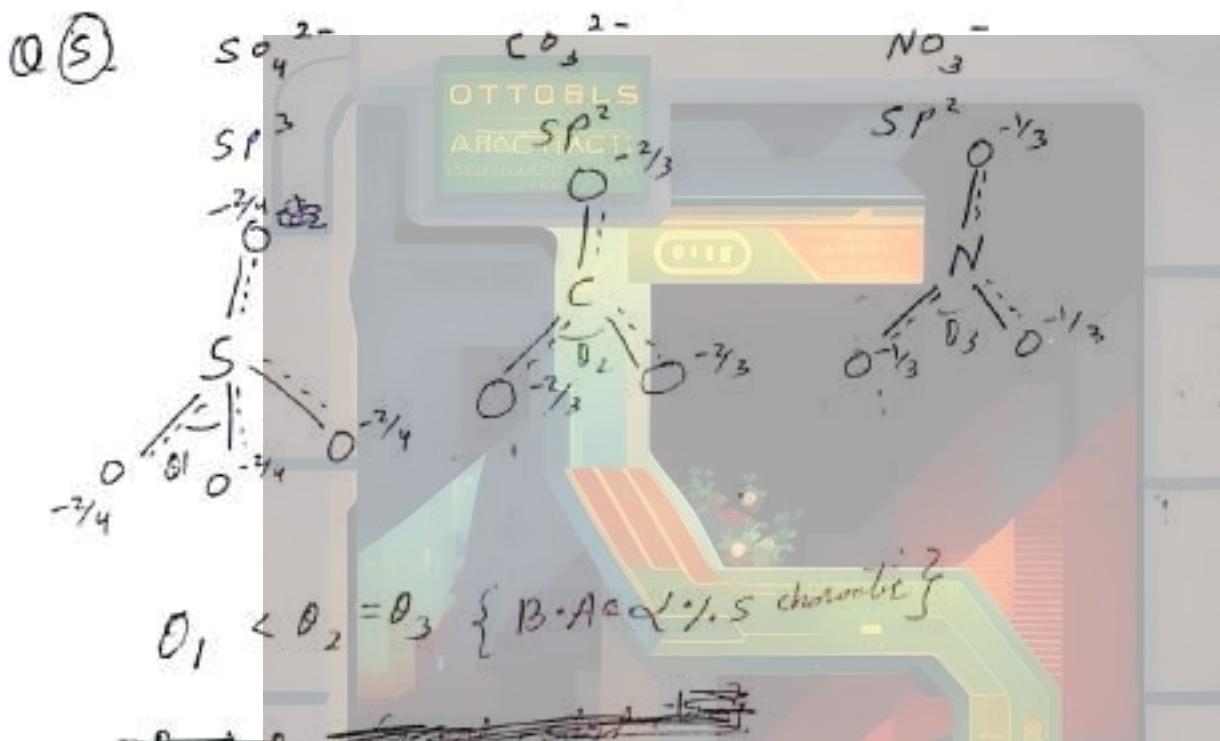
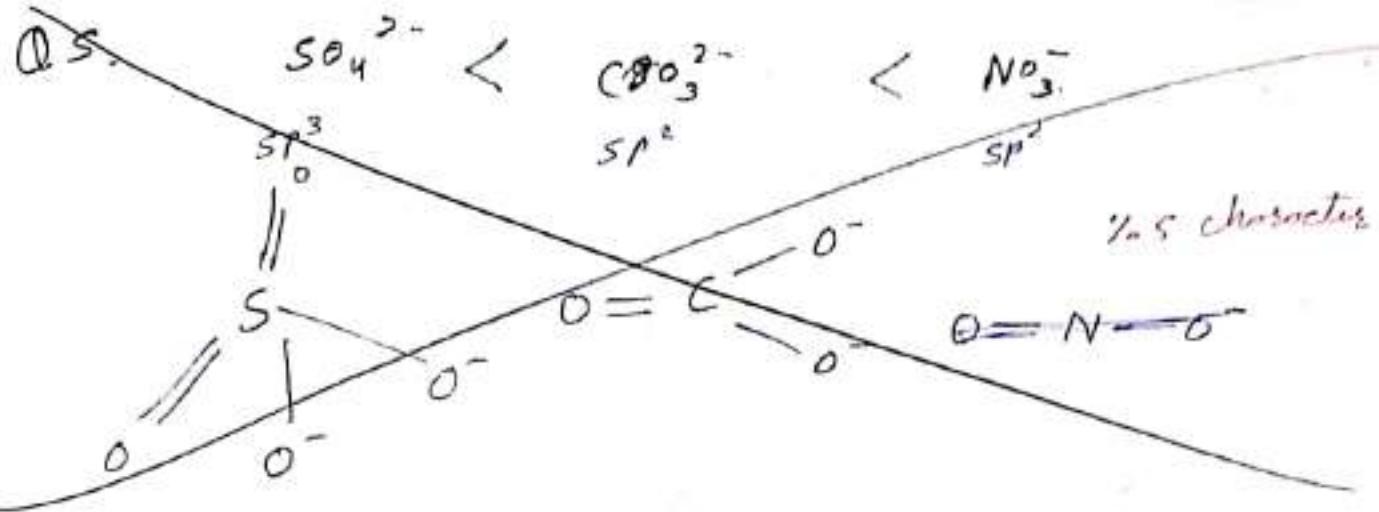


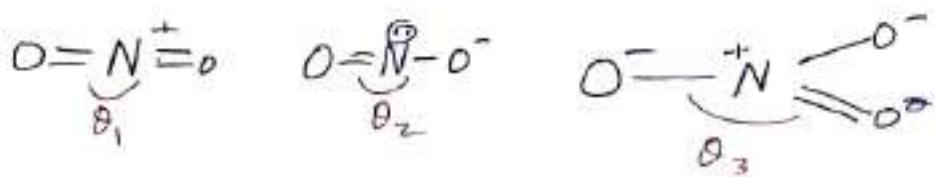
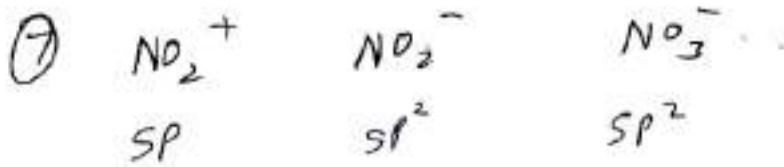
④



EN of central atom

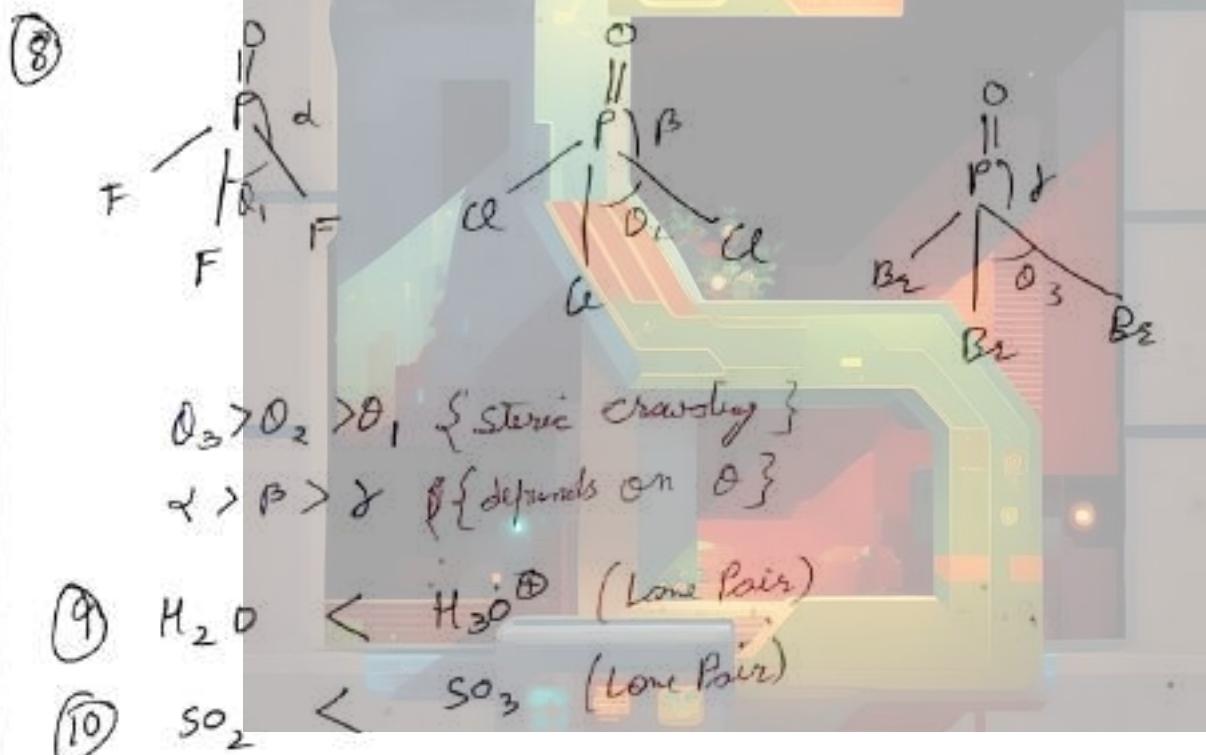
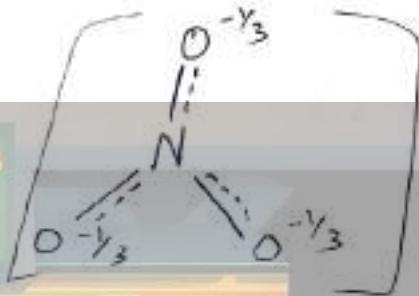
$\text{EN}(\text{s}) > \text{EN}(\text{se})$

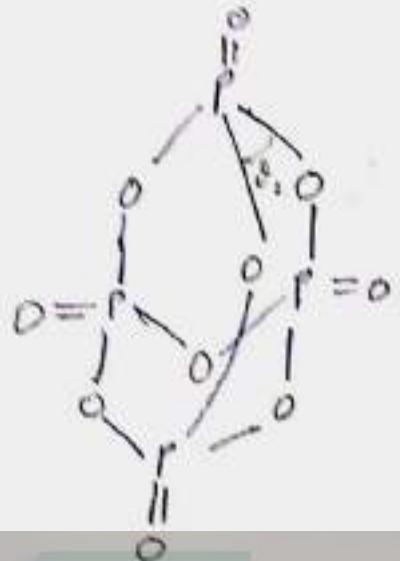
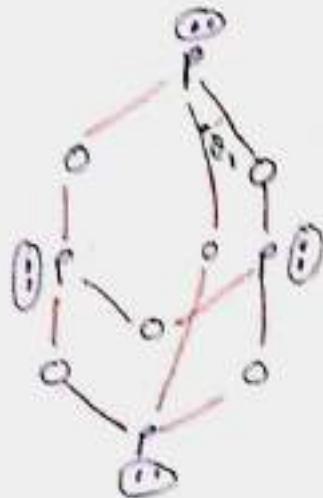
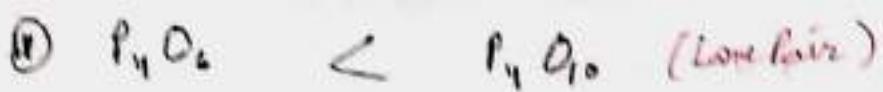




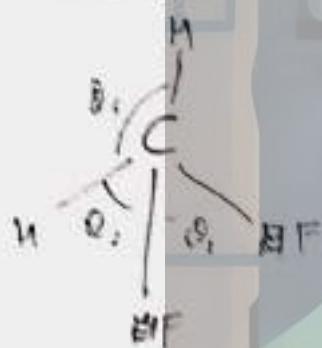
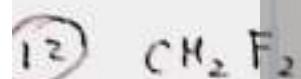
$$\boxed{\theta_1 > \theta_3 > \theta_2}$$

↓

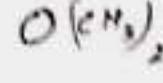
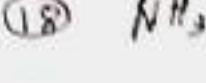
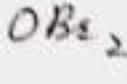
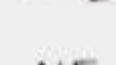
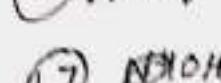
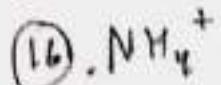
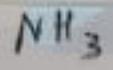
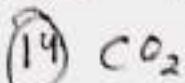
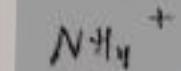
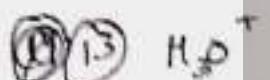


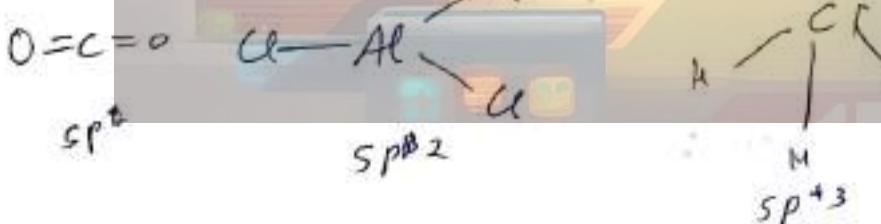
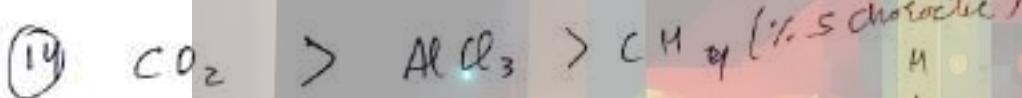
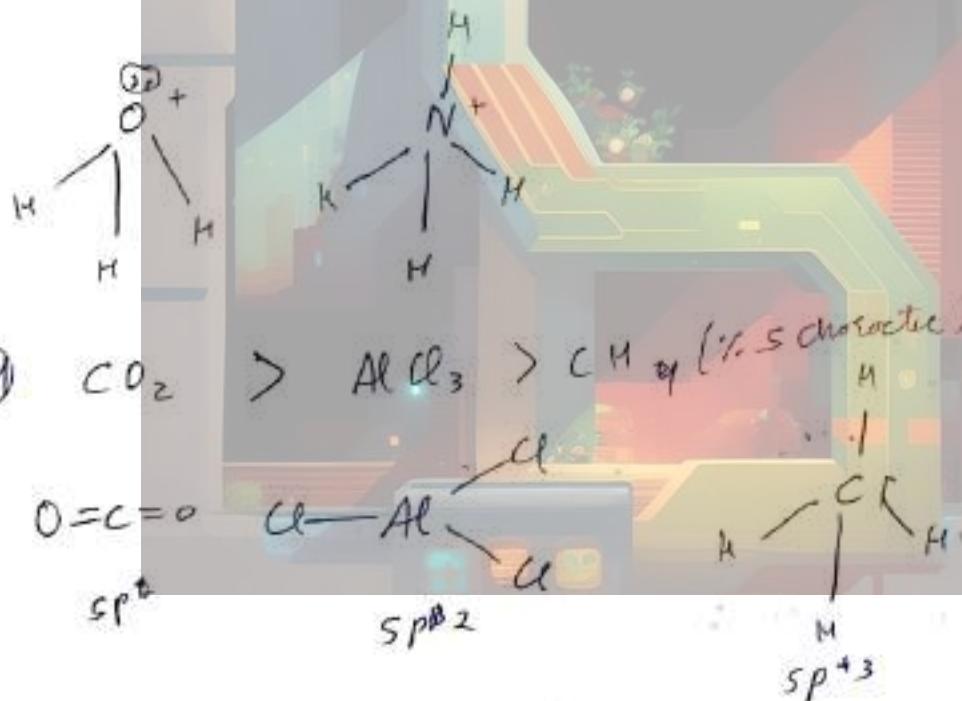
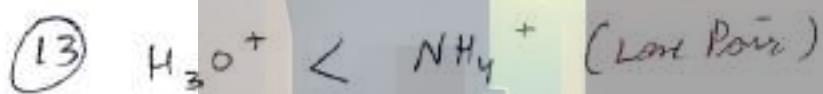
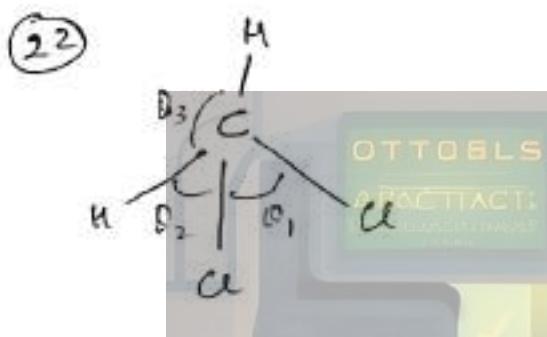
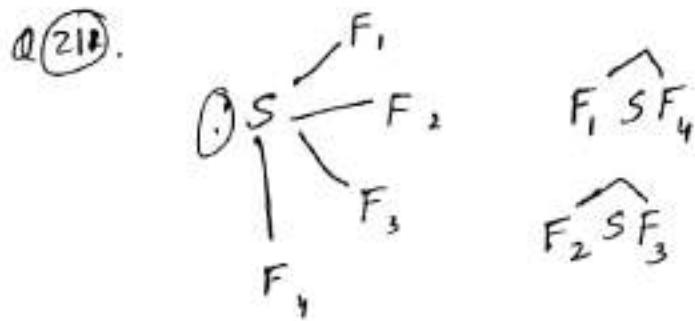
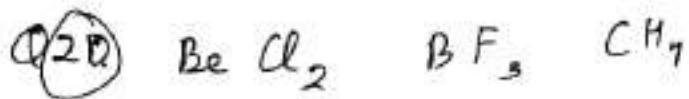


$\angle P_{\text{dipole}} > \text{D-B-F symm}$

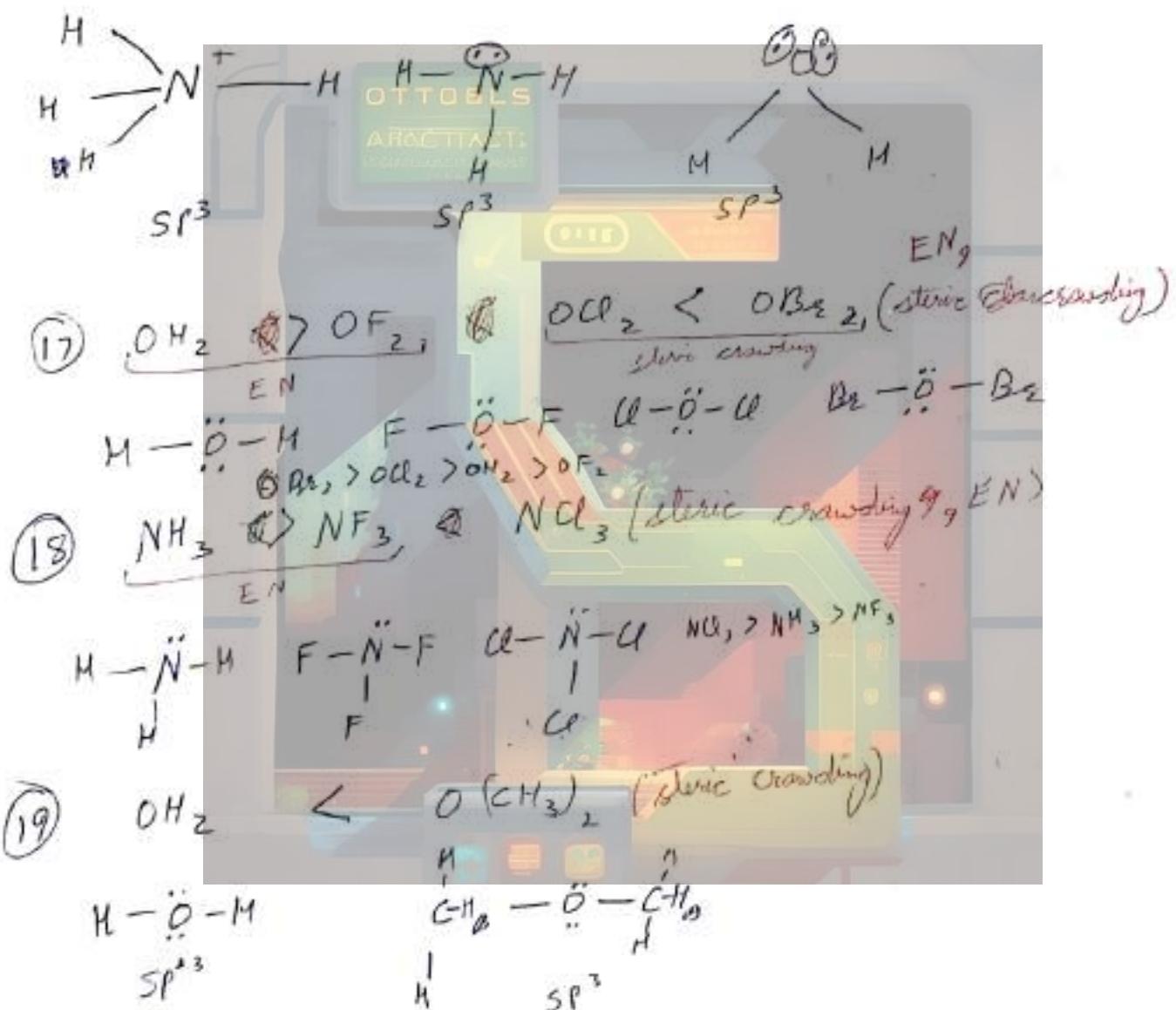
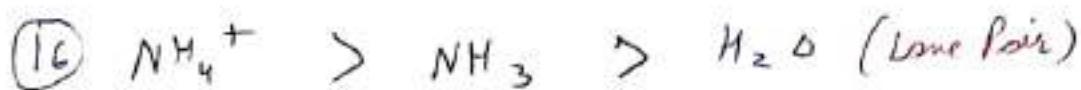
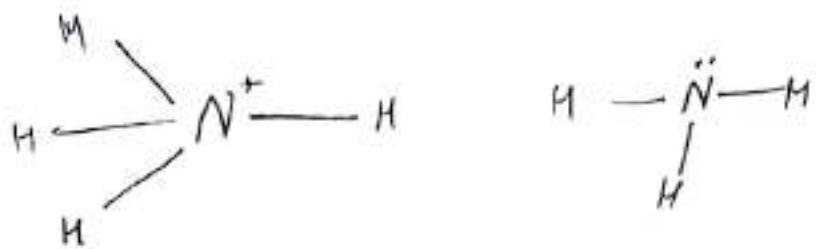
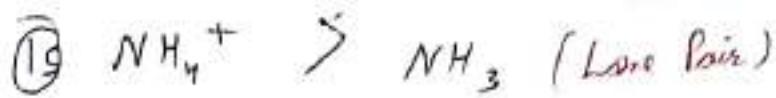


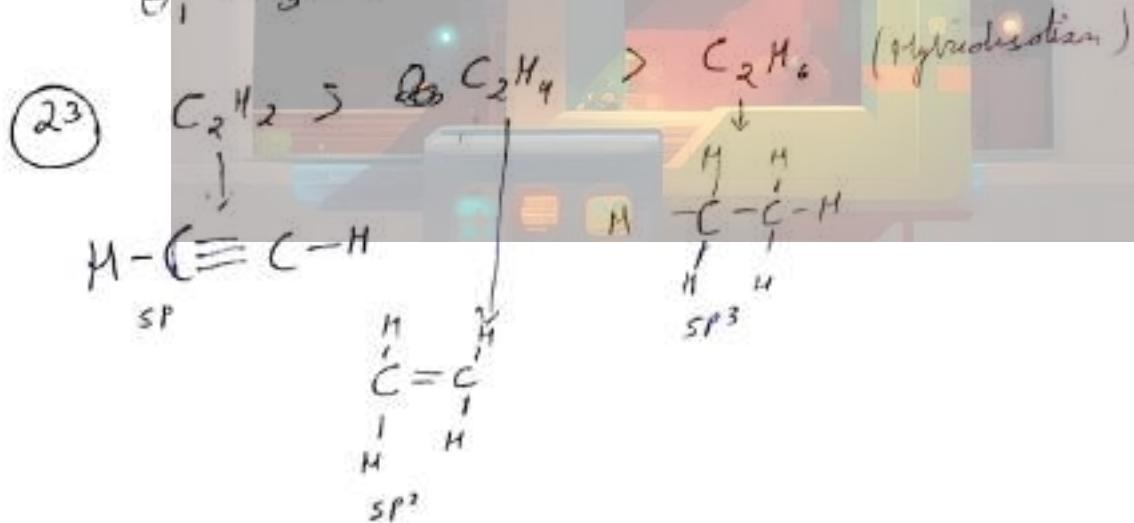
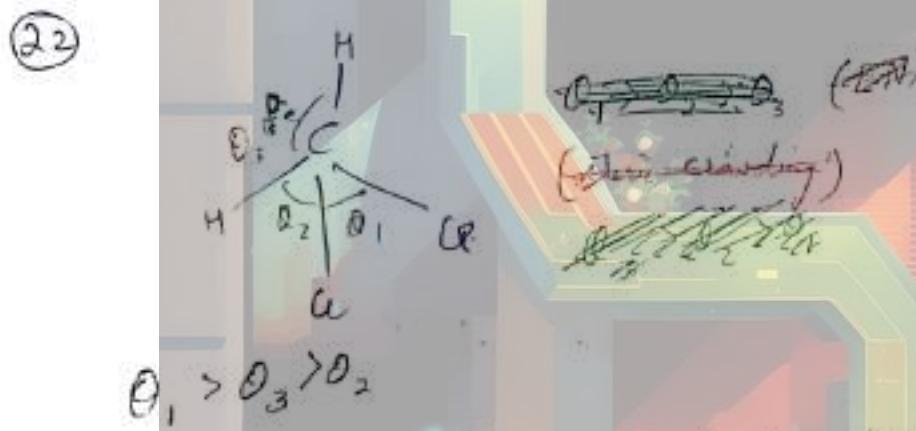
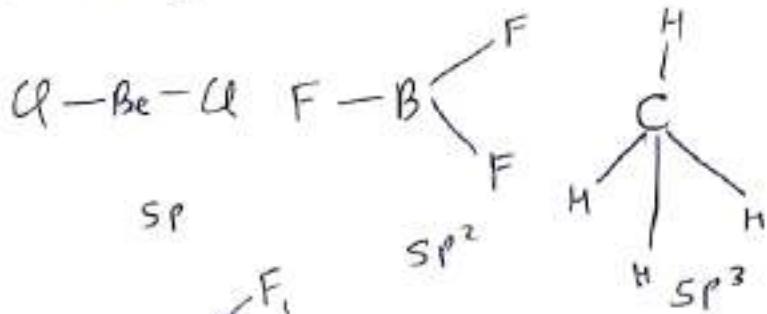
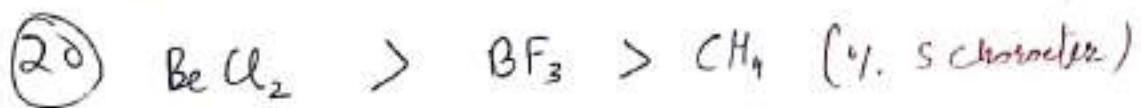
$O_3 > O_2 > O_1$ (E.N. variation)





(38)

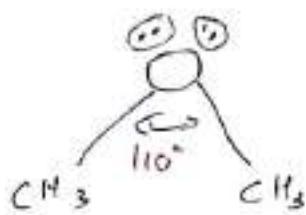




(40)

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

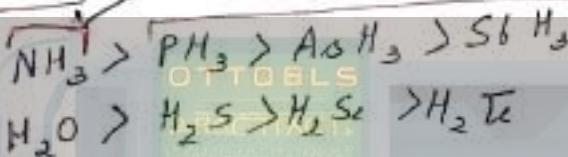
(25) $O(CH_3)_2 < OCl_2 < OBr_2 < OI_2$ (Exception)



Dobro's Rule :-

Hybridized fail

B.A.



$PH_3 \rightarrow 94^\circ$

$AsH_3 \rightarrow 92.5^\circ$

$SbH_3 \rightarrow 91^\circ$

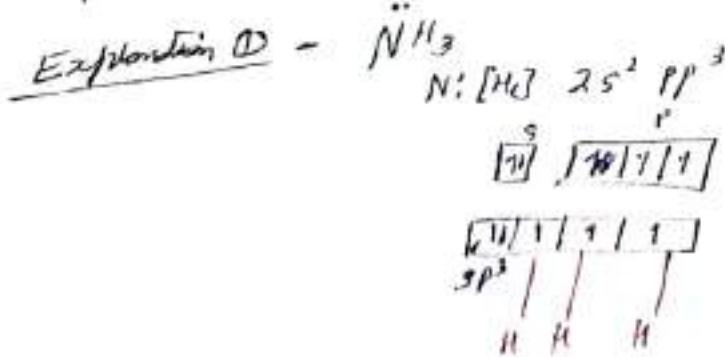
$H_2S \rightarrow 92^\circ$

$H_2Se \rightarrow 91^\circ$

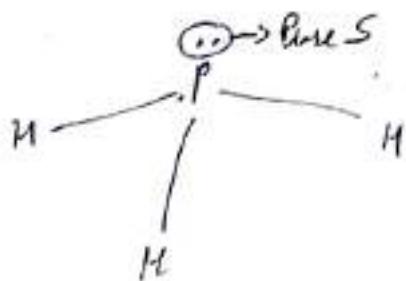
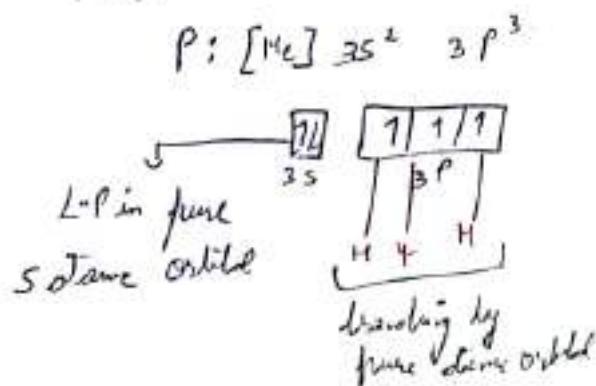
$H_2Te \rightarrow 90^\circ.5^\circ$

Angles are around 90° means pure atomic or participating more than hybrid orbitals.

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



pH_3



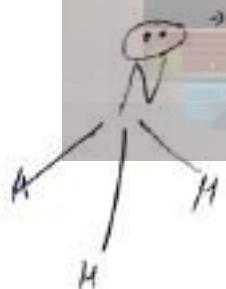
Explanation - 2 :- The cause of no hybridisation in this molecule can be explained as follows.

→ In case of PH_3 , the energy required in hybridisation is about 600 kJ/mol which is not compensated by energy released from the bond formation using hybrid orbitals.

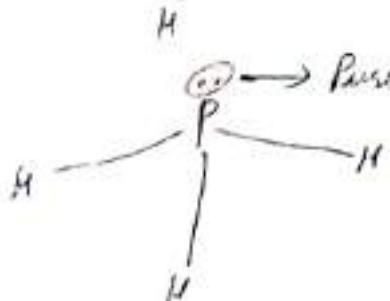
Q1. Ammonia is strong Lewis base as compared to Phosphine (PH_3)

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

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

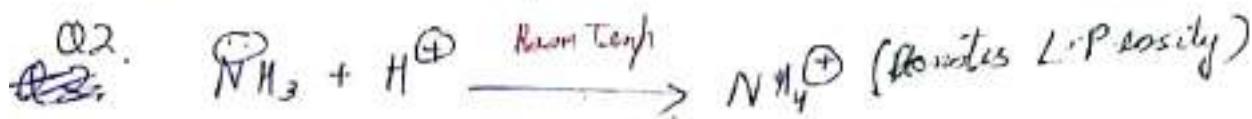


$\rightarrow \text{sp}^3 \text{ orbital} \Rightarrow 25\% \text{ s character} \Rightarrow$ Attraction from nucleus is less
 \downarrow
 e^- remove easily



\rightarrow 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 denote lone pair easily & as $\text{NH}_3 \text{L.P}$
 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 the sp^3 hybridized orbital which is more directional & more reactive.

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

⇒ Types of Hybridisation.

Equivalent Hybridisation

- no L.P. on central atom
- surrounding atoms must be same
- All bond angles are same

Eg $\text{CH}_4, \text{CCl}_4, \text{BF}_3$

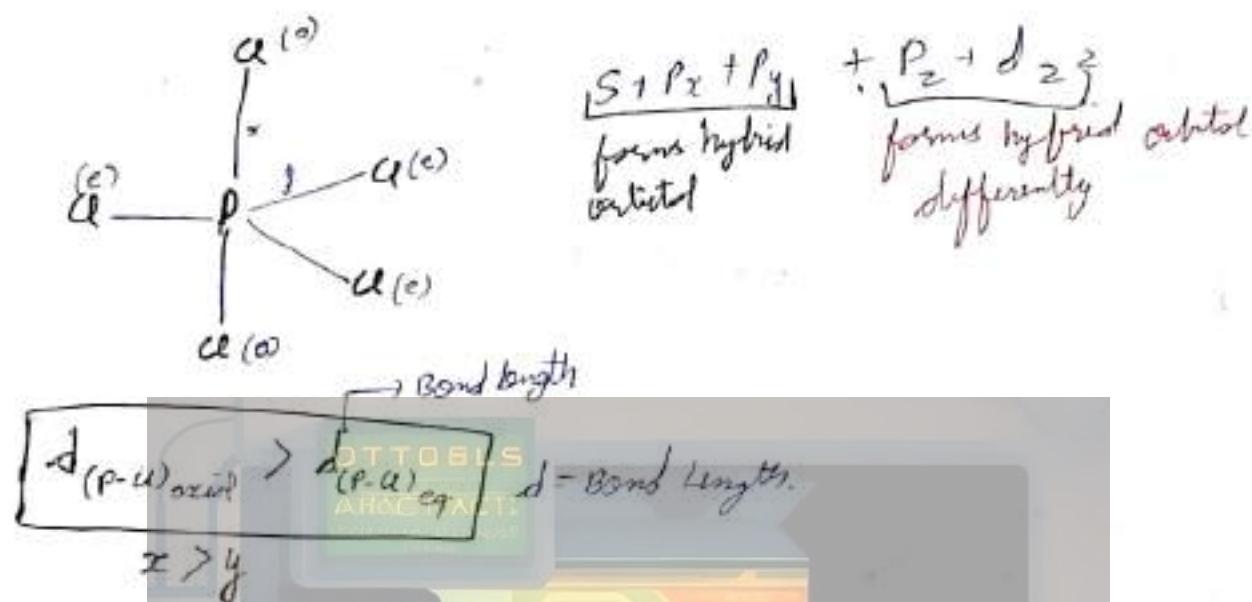
Non Equivalent Hybridisation

- It. 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 PCl_5 , it is observed that there are two longer (weaker) axial bonds and 3 shorter (stronger) equatorial bonds are present.

→ Experimentally it is concluded that all hybrid orbitals are not equal in PCl_5 .

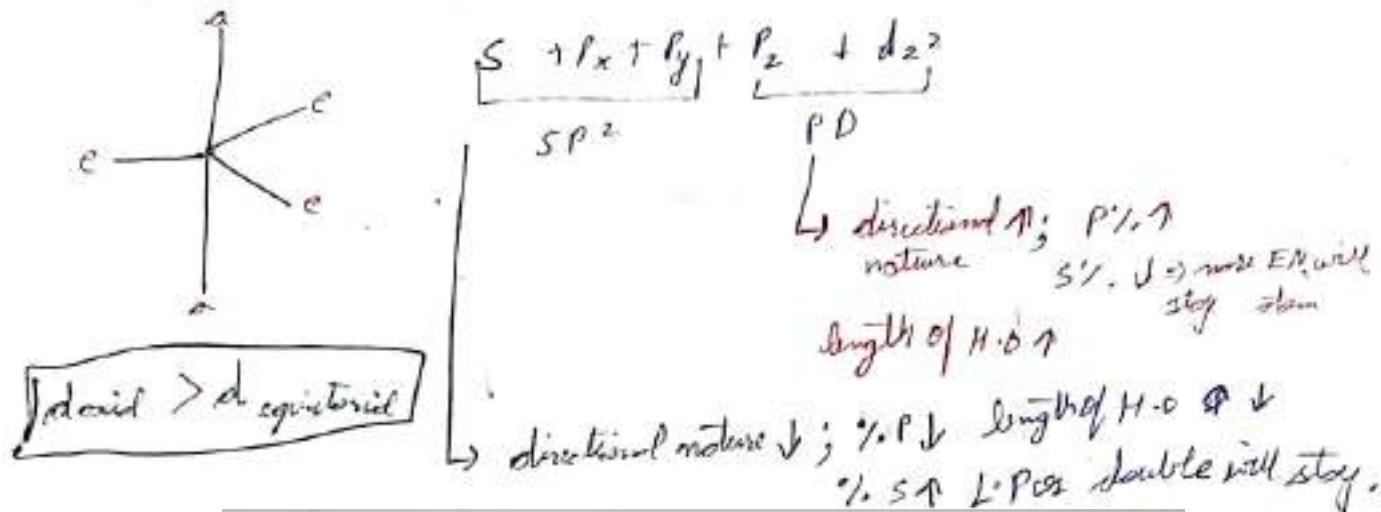
→ Equatorial bonds are formed by sp^2 hybrid orbitals and axial bonds are formed by p_z PD hybrid orbitals of central atom.

According to Bent's rule.

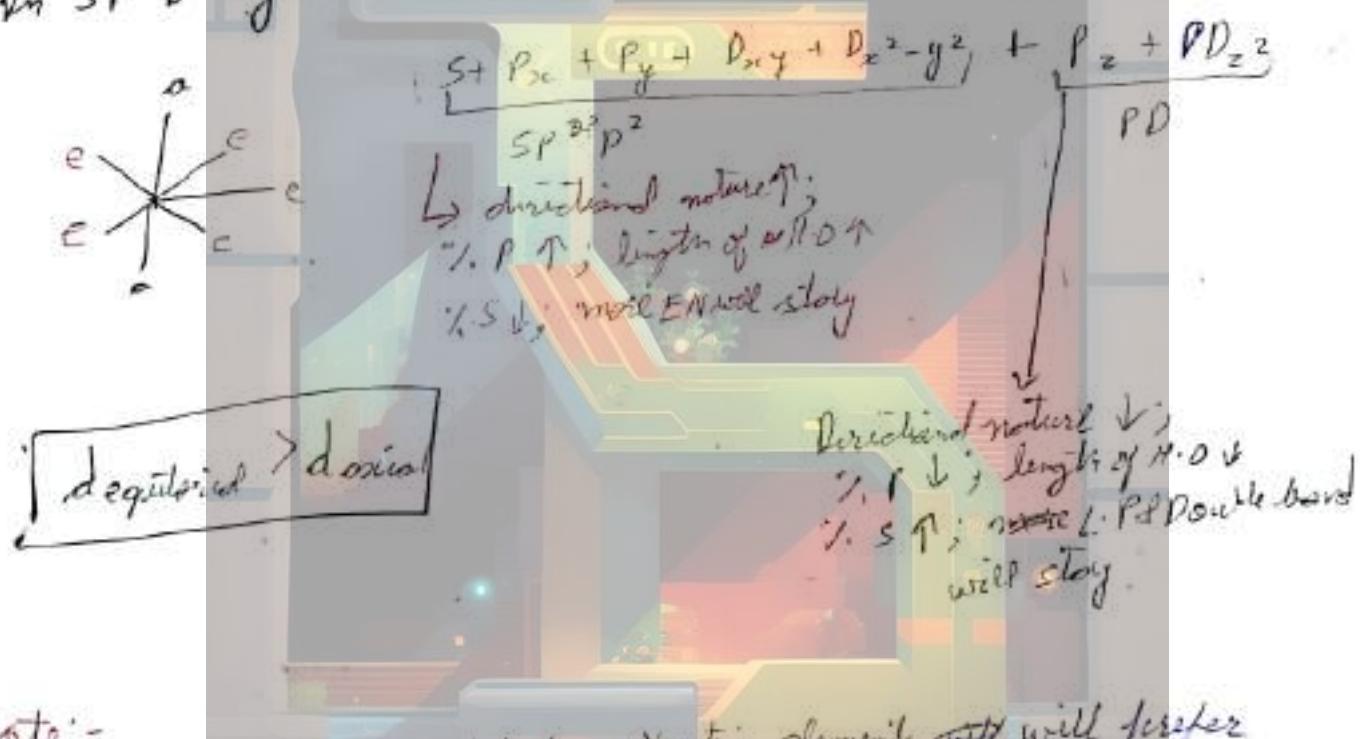
→ More Electron Negative element (surrounding atom) prefer to stay in the hybrid orbital of central atom, which is having less "s character" far from nucleus.

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

In SP^3D hybridisation (TBP)



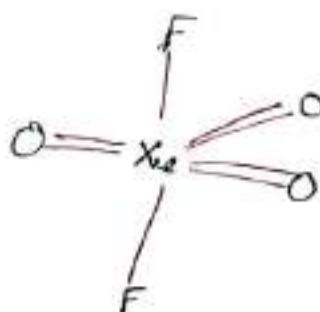
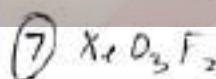
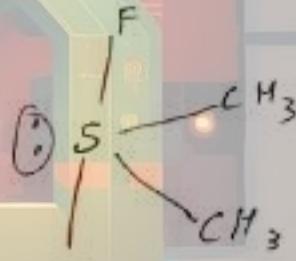
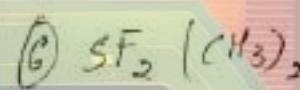
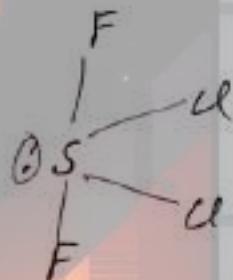
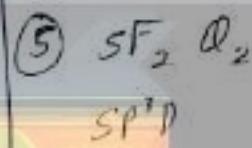
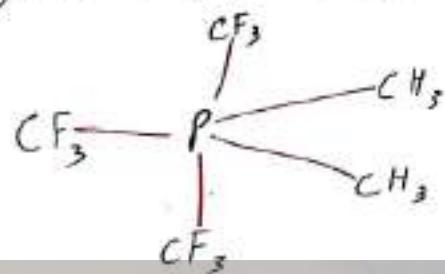
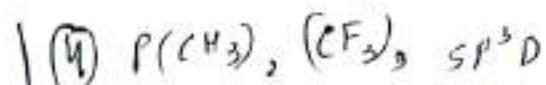
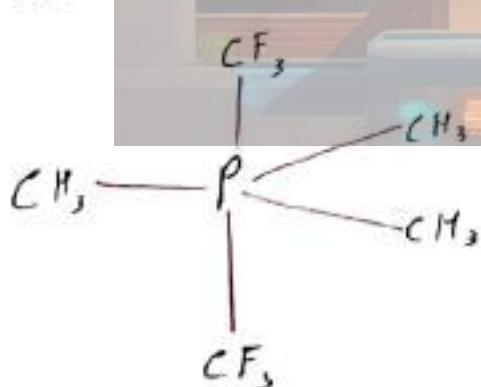
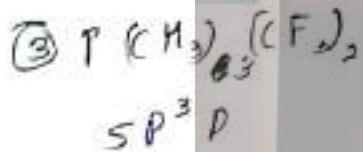
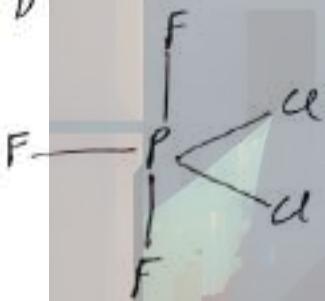
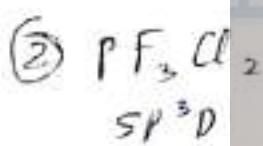
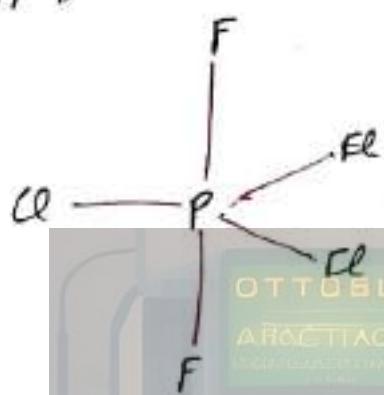
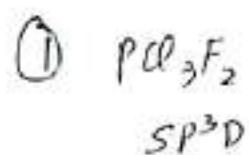
In SP^3D^3 hybridisation (PBP)

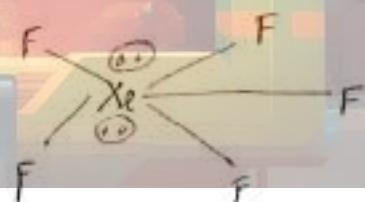
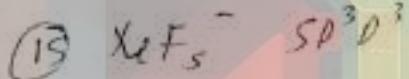
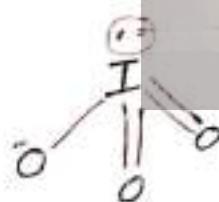
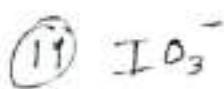
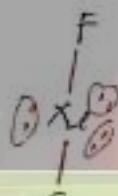
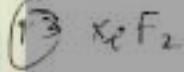
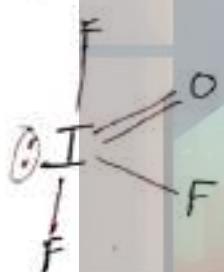
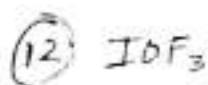
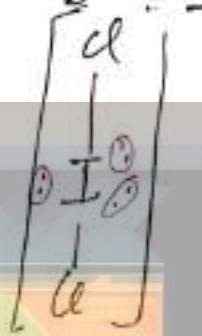
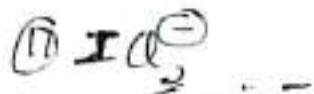
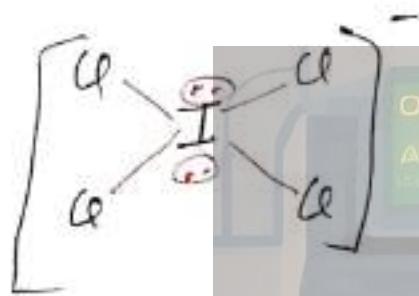
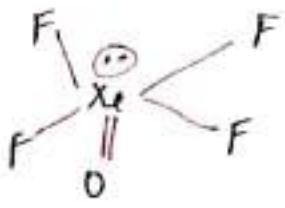
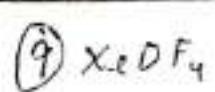
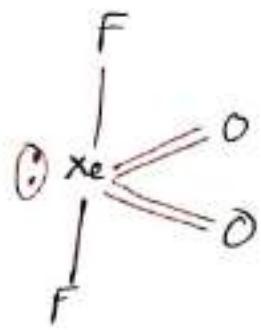
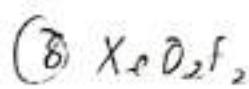


Note:-

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

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

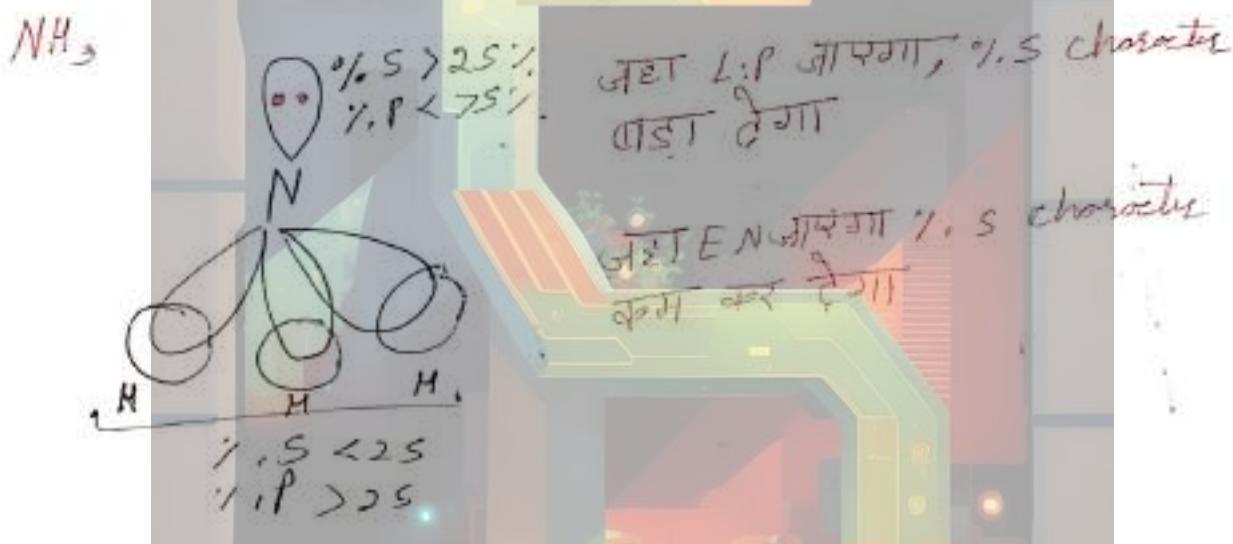




17

Alternative Statements of Bent's Rule:-

- More E-N atom not only prefer to stay in the hybrid orbital of centred atom which is having less %s character but also it decreases %s character to attached hybrid orbital of centred atom which depends upon circumstances.
- L-P not only to prefer to stay in the hybrid orbital of centred atom which is having more %s character but also it increases %s character to attached hybrid orbital of centred atom which depends upon circumstances.



How To Solve Questions of Bond Length.

Rule 1 - If shell no. of central atom or surrounding atom are different.

Shell no. ↑ Bond Length ↑

e.g. $H-F \quad H-Cl \quad H-Br = H-I$

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

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

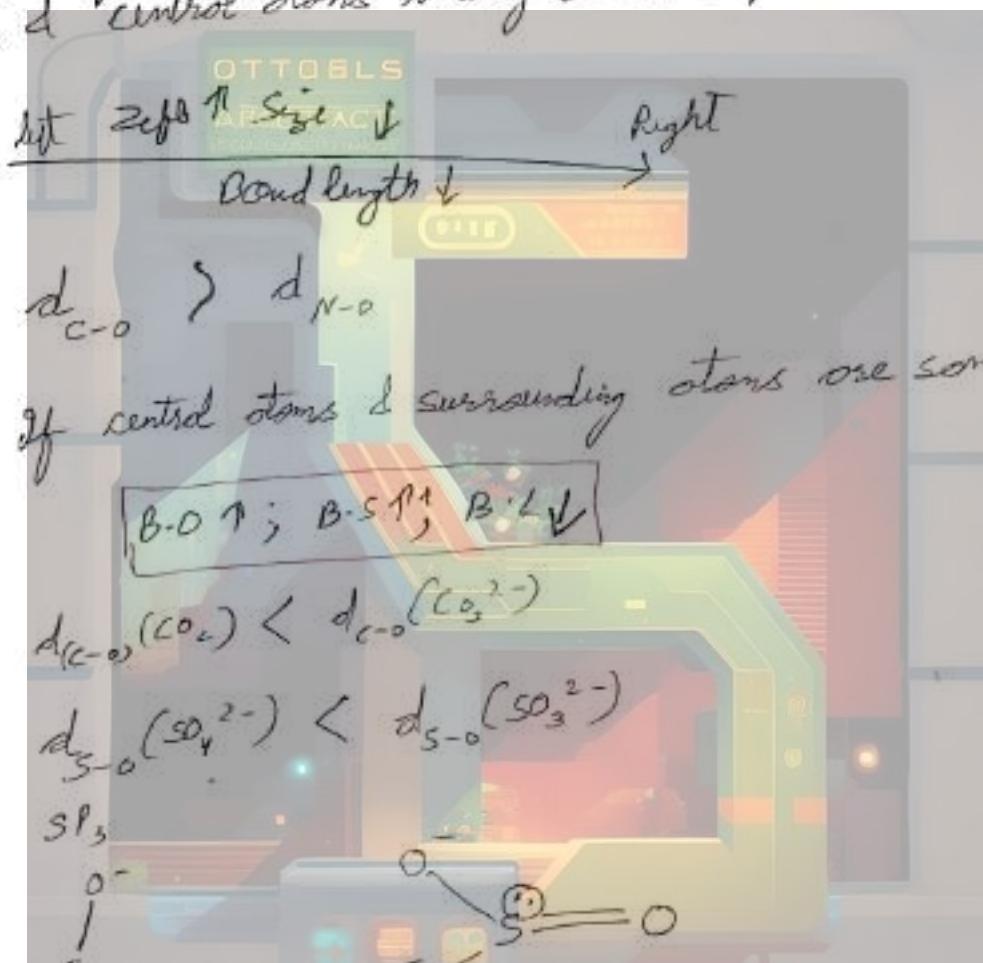
$$\text{Eg. } d_{(P-O)} > d_{(S-O)} > d_{(C-O)}$$

$$r_p > r_s > r_c$$

$$\text{Eg ③. } d_{(P-F)} < d_{(P-C)}$$

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

Rule ② → If surrounding atoms are same, central atoms are different & control atoms belong to same period.



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

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



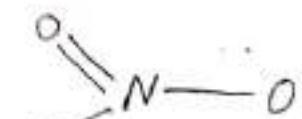
$$BO = \frac{5}{4}$$

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

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



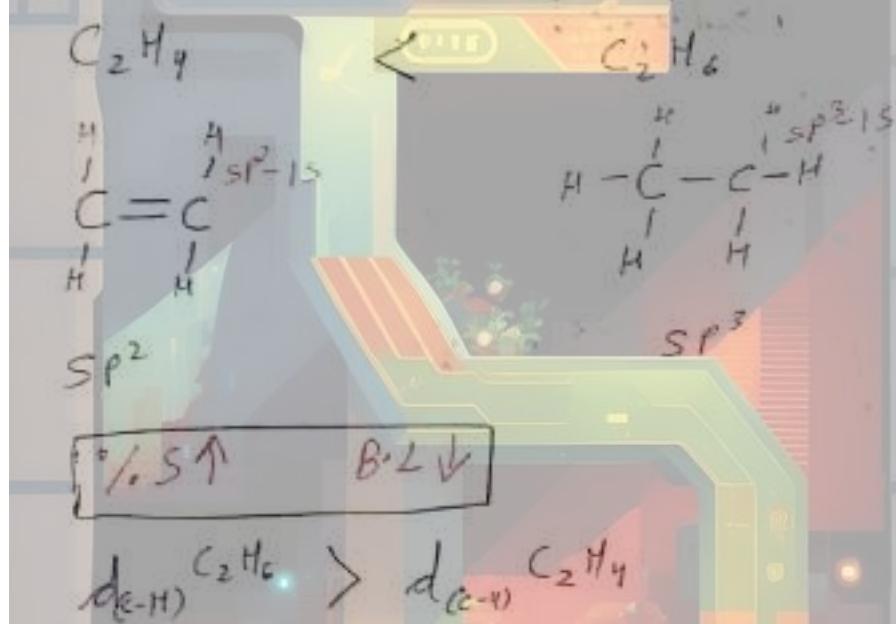
$$B.O = \frac{3}{2}$$



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

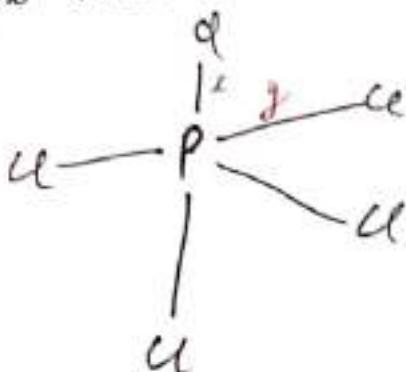
Rule-④ → If central atom, surrounding atom and Bond order is same then Bond Length depends on hybridisation.

Eg①.



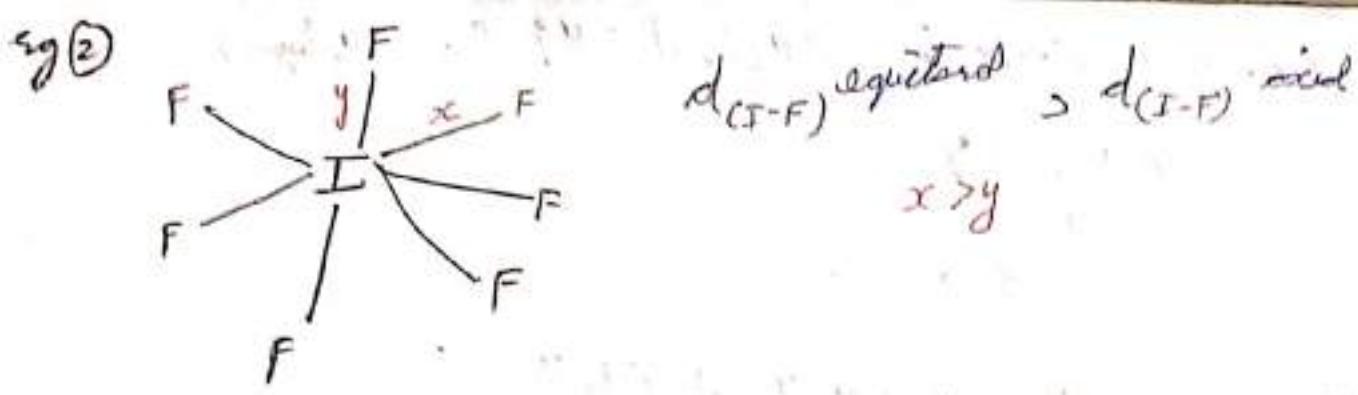
Rule-⑤ → If central atom, surrounding atoms, Bond orders and hybridisation are same then Bond length is decided by bent's rule.

Eg①

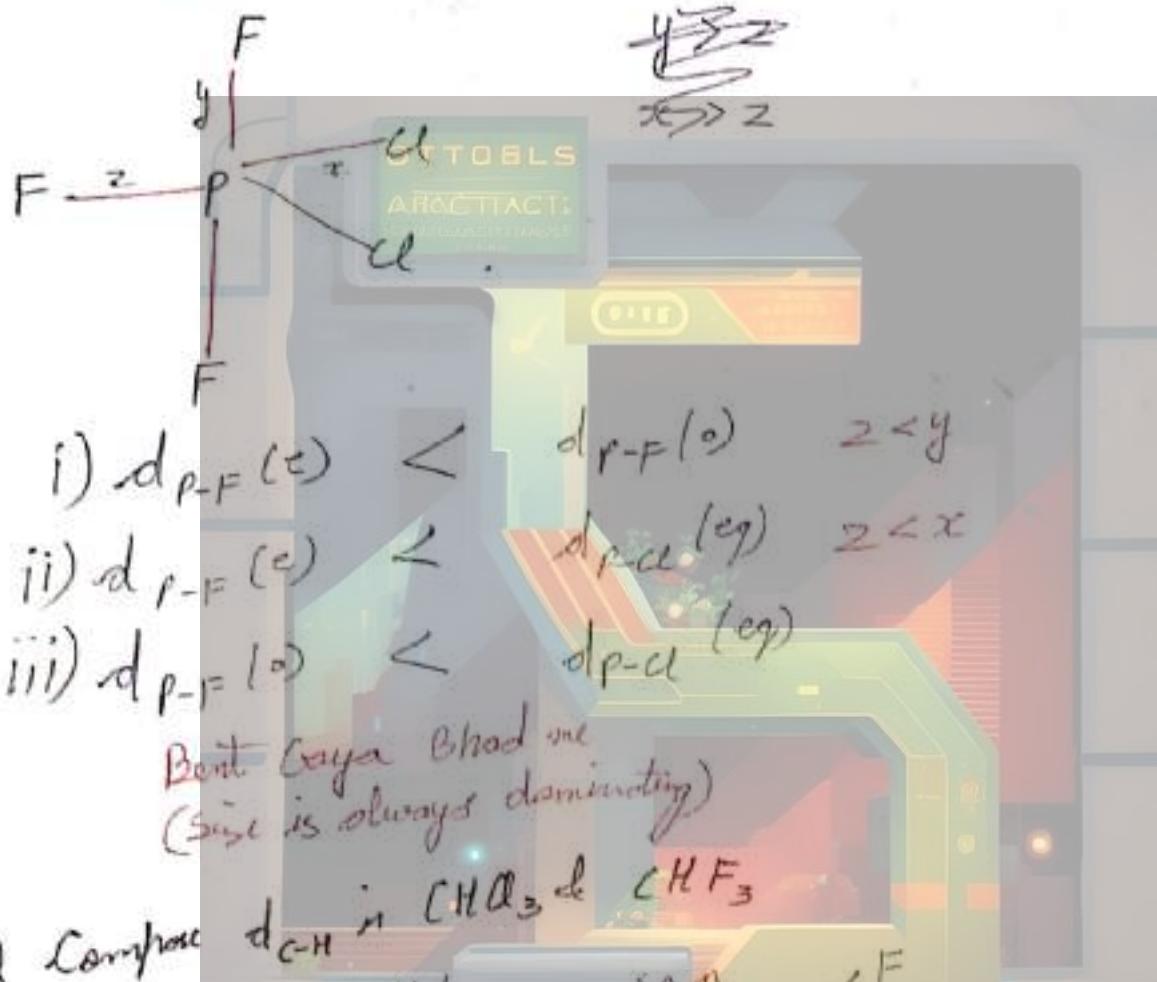


$$d_{(\text{p-axial})} > d_{(\text{e-equatorial})}$$

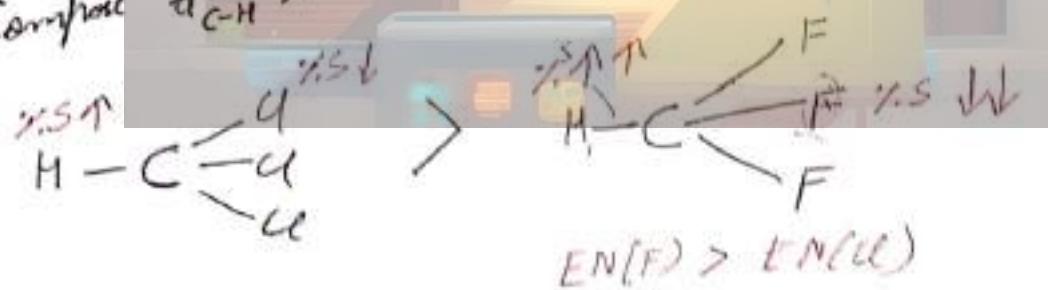
$$x > y$$



Q PF_3Cl_2

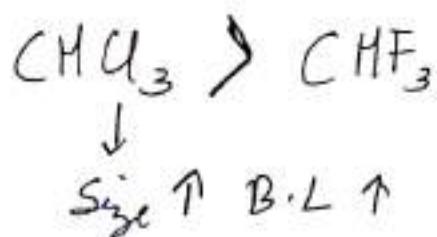


Q Compare d_{C-H} in $CHCl_3$ & CHF_3

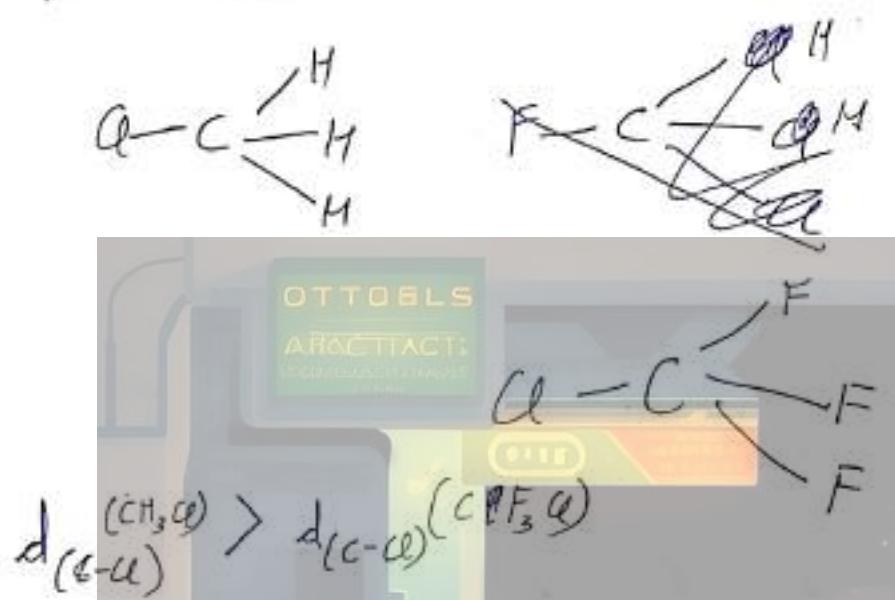


~~CHCl₃~~ $y.s \uparrow$ $y.s \downarrow$ $y.s \uparrow$ $y.s \downarrow$, $EN(Cl) > EN(F)$

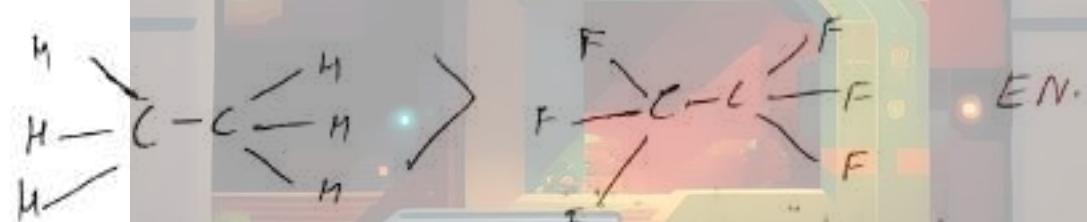
Q2. Compare C-Cl in CHCl_3 & CHF_3 (\rightarrow Hologram)



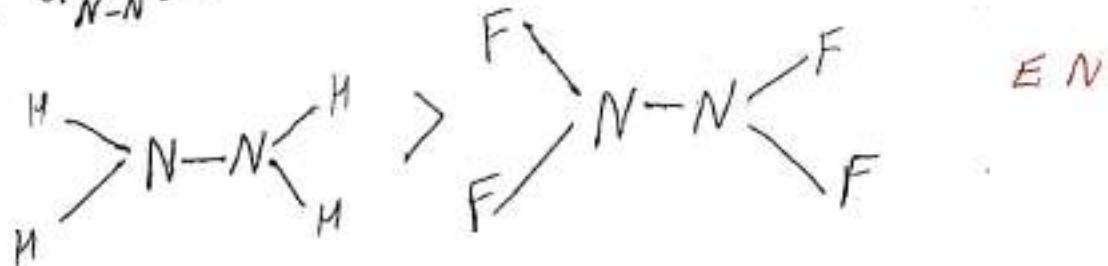
Q3. compare $d_{\text{C-Cl}}$ in CH_3Cl & CF_3Cl

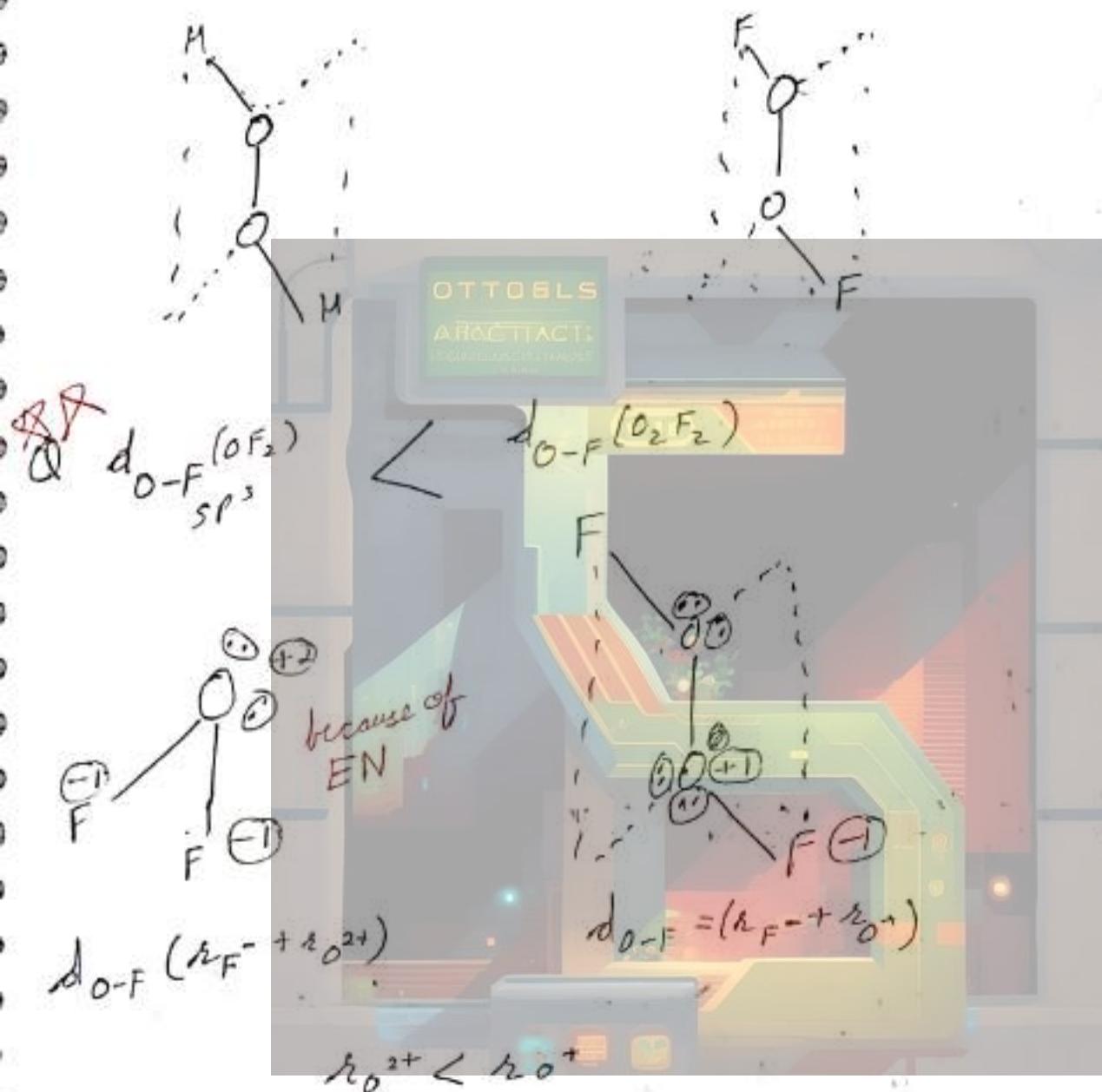
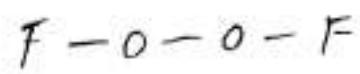
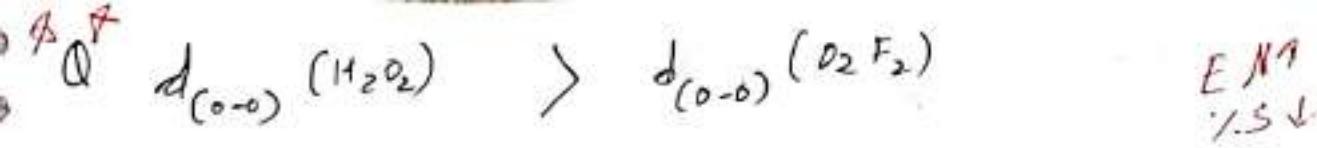


Q4. Compare C-C in C_2H_6 & C_2F_6

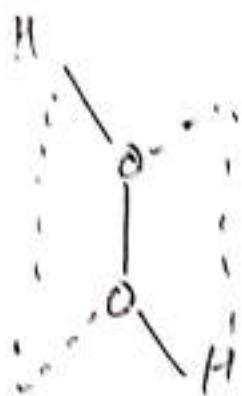


Q5. $d_{N-N}(\text{N}_2\text{H}_4)$ & $d_{N-N}(\text{N}_2\text{F}_4)$

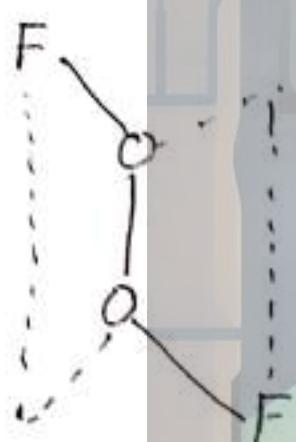




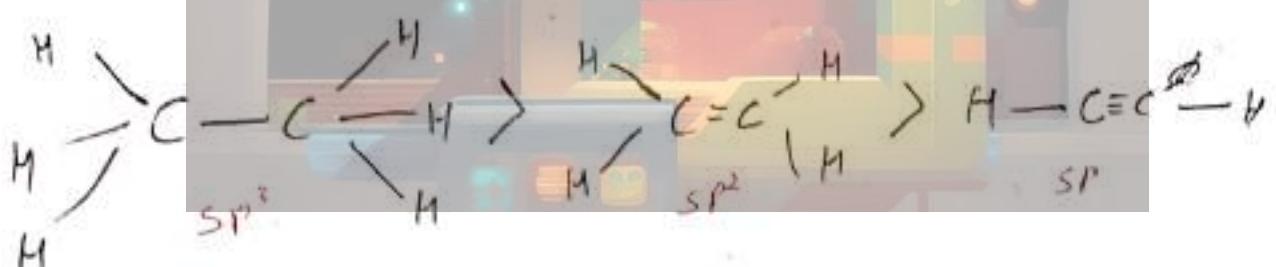
$$\text{Q } d_{(O-O)}(H_2O_2) > d_{O-O}(O_2)$$



$$\text{Q } d_{(O-O)}(O_2F_2)$$

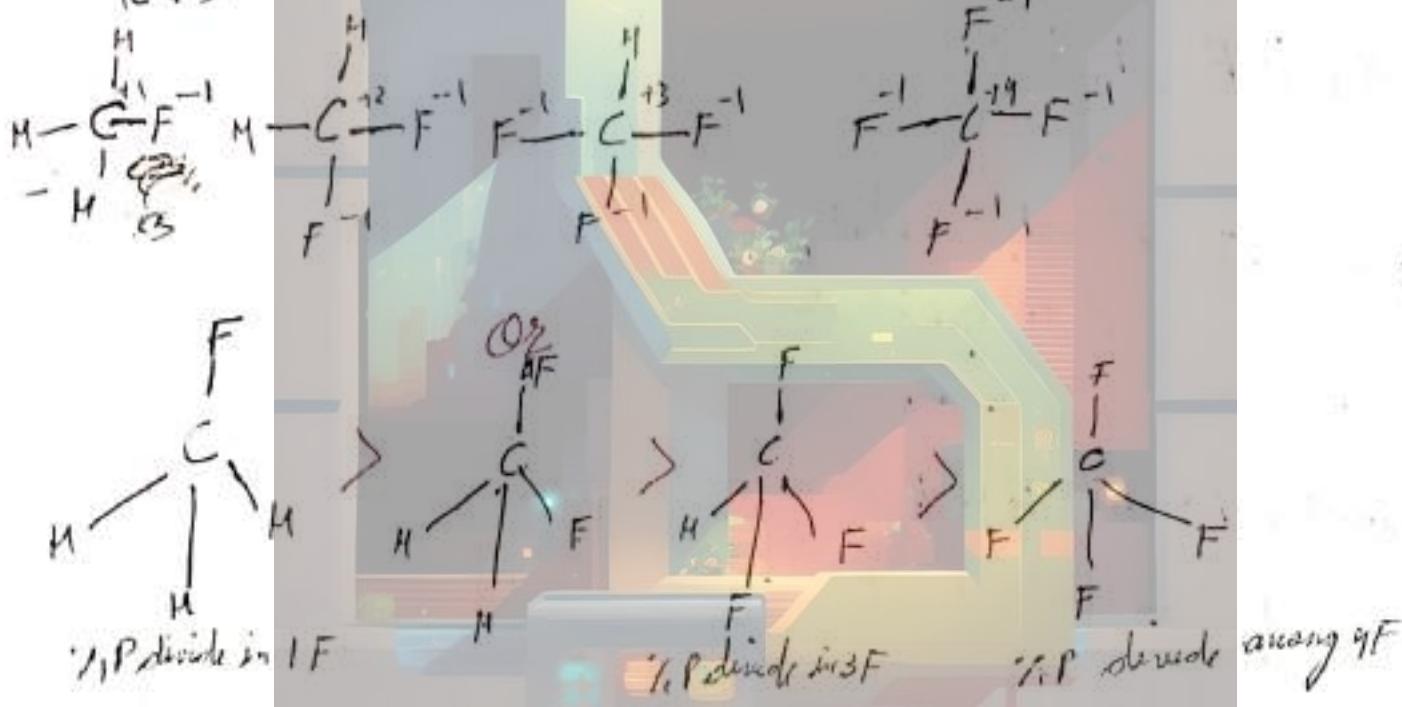
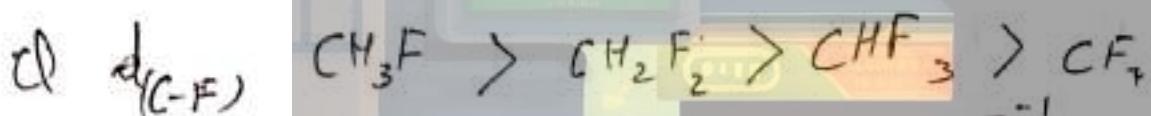
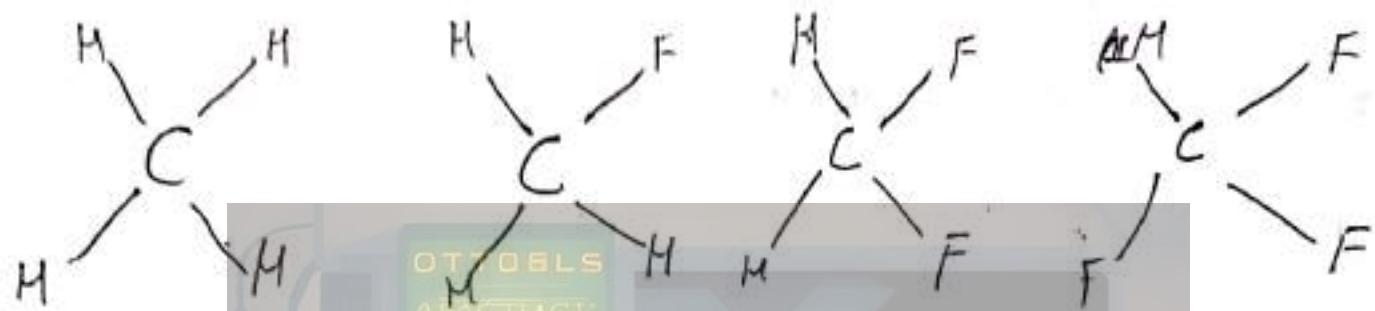
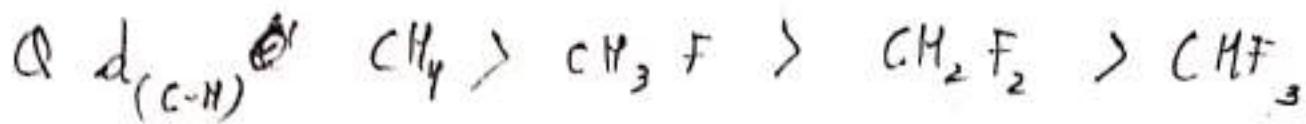
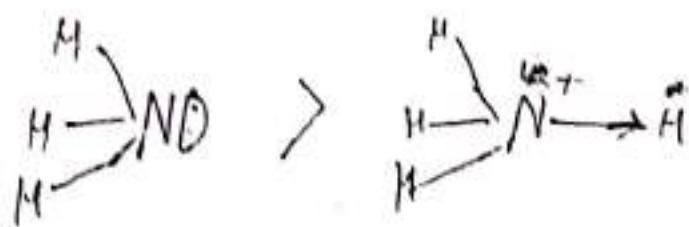


$$\text{Q } (C-C)_m(C_2H_6), C_2H_4, C_2H_2$$



$$\text{Q } d_{(S-O)} SO_2 \text{ & } SO_3$$

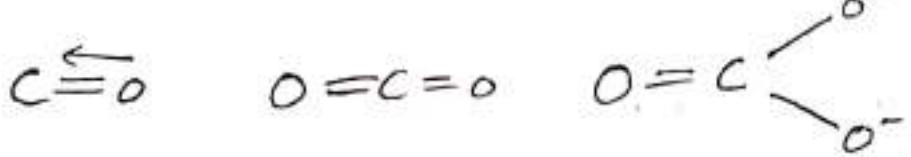




Bond Energy \rightarrow Bond Energy is the energy released when one mole of some type of bond is formed.

\rightarrow Bond Energy is the energy supplied to break 1 mole of some type of bond.

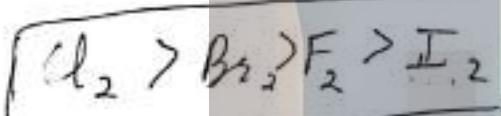
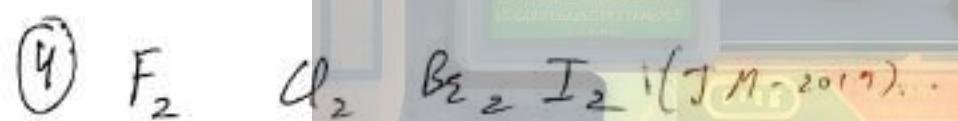
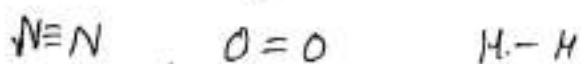
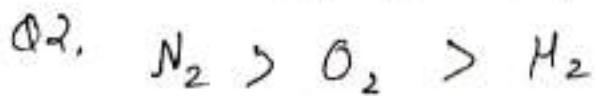
$B.O \uparrow$	$B.E \uparrow$
----------------	----------------



$$\text{B.O} = 3$$

$$\text{B.O} = 2$$

$$\text{B.O} = \gamma_3 = 1.33$$

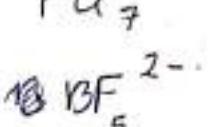
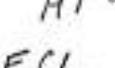
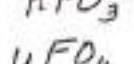
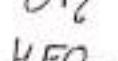
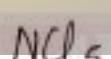
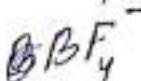
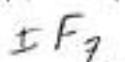
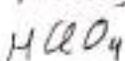
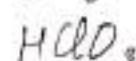
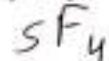
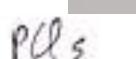


Note - $\sqrt{s + p_x^2 + p_y^2 + d_{x^2-y^2}}$
d sp^2 hybridization

Structure :- Square Planes

Molecules Do not Exist

① Lack of 2D molecular orbital :- not Exist



Note - Second Period elements cannot expand their octet as they do not have 2d subshell or orbital.

② Non Existence due to steric crowding:-

exist

PF_5

SF_4

SF_6
 IF_7

not exist

PI_{Scg}

SCl_4 , SBr_4 , SE_4

SCl_{Sc} , SBr_{Sc} , SE_{Sc}

ICl_7 , IBr_7

③ VBT :- (exist for 2nd period)

exist

PF_5

SF_4
 SF_6

IF_3

IF_{Sc}

IF_7

not exist

PF_4

SF_3

SF_5

IF_2

IF_4

IF_6

} odd e⁻ species

④

④ Absence of D-orbital contraction:-

not exist

exist

PF_5

SF_4

SF_6

IF_3

IF_7

XeF_3

XeF_4

PH_5

SH_4

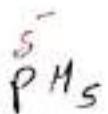
SH_6

IH_3

IH_7

XeH_2

XeH_4



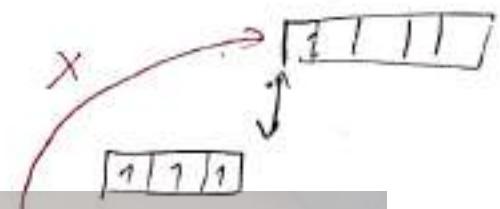
P:

1	1	1	1
---	---	---	---

1	1	1	1
---	---	---	---

1	1	1	1	1	1
---	---	---	---	---	---

$SP^3 D$



The S^+ charge reduce size,
so as the orbitals come
closer it requires less
energy to hybridise orbitals

S^- size increase size,
as the orbitals are far'd
requires more energy to
hybridise the orbitals

But hydrogen don't provide
enough high energy to
hybridise the orbitals.

→ Such species do not exist in which all surrounding
atoms are hydrogen (Electron Negativity ≤ 2.1) and
 σ -orbital of central atom is participating in hybridisation

Dipole Moment :-
2 oppositely charged

→ Due to Difference in E.N., two poles are created in covalent bond



$$E.N_{(A)} > E.N_{(B)}$$

→ The two charged poles act as electric poles and its degree of polarity is expressed as Dipole moment.

→ It is the product of ^{OPPOSITE} charge accumulated on element & interatomic distance (Bond length)

→ It is a vector quantity whose magnitude is:

$$\mu = q \times d$$

μ - Dipole moment
q - Charge (Coulomb)
d - Interatomic Distance/Bond length

Dipole moment.

& direction from Positive to Negative

→ S.I unit - C.m (coulomb meter)

C.G.S - CGS - cm

- Debyē (D)

$$1 D = 3.33 \times 10^{-30} \text{ esu m (C.m)}$$

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

→ For a molecule, Dipole moment is a net vector addition of all individual bond dipole moment.

$$\mu_{\text{resultant}} = \sqrt{\mu_1^2 + \mu_2^2 + 2\mu_1\mu_2 \cos\theta}$$

μ_1 & μ_2 - Dipole moment

θ Bond angle b/w the dipole moment.

Application

- ① To calculate % I.C in covalent bond:
- $$\% \text{ I.C} = \frac{\mu_{\text{observed}}}{\mu_{\text{calculated}}} \times 100 \quad \rightarrow \text{Experimental}$$

for single boron halide molecule

$$\% \text{ I.C} = \frac{\mu_{\text{observed}}}{1.6 \times 10^{-19} \text{ C m}} \times 100$$

Q1. Calculate % I.C in HCl molecule if dipole moment $\mu = 1.03 \text{ D}$

$\& B \cdot L = 1.56 \text{ pm}$
 $= 1.56 \text{ } \text{\AA}$
 $= 1.56 \times 10^{-12} \text{ m}$

$\% \text{ I.C} = \frac{3.4299 \times 10^{-30}}{1.56 \times 1.6 \times 10^{-19} \times 10^{-30}} \times 100$

$= \frac{3.4299}{10^9 \times 2.496}$

$= \frac{3.4299}{2.496 \times 10^7}$

$= \frac{3.4299}{2.496 \times 10^9}$

$= \frac{3.4299}{2.496} \times 10^{-7}$

$\% \text{ I.C} = \frac{3.4299}{1.56 \times 10^{-12} \times 1.6 \times 10^{-19}} \times 10^{-2}$

$= 131.741 \%$

Q Calculate B.L of HBr molecule if 1. I.C is 11.5%
2 Dipole moment is $2.6 \times 10^{-30} \text{ C m}$

$$11.5 = \frac{2.6 \times 10^{-30} \times 10^2}{1.6 \times 10^{-19} \times d}$$

$$d = \frac{2.6}{1.6 \times 11.5} \times 10^9$$

$$d = 0.141 \times 10^9$$

$$d (\text{in pm}) = 0.141 \times 10^9 \times 10^{-12}$$

$$= 0.141 \times 10^{-3}$$

$$= 1.41 \times 10^{-3} \text{ pm}$$

$$d = \frac{2.6}{1.6 \times 11.5} \times 10^{-9}$$

$$\approx 1.41 \times 10^{-9}$$

$$= 0.141 \times 10^{-9} \times 10^{12}$$

$$= 0.141 \times 10^{-3}$$

$$\boxed{d = 141.3 \text{ pm}}$$

(2) Polarity of molecule species.

a) Polarity of bond

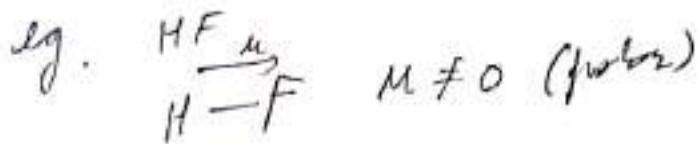
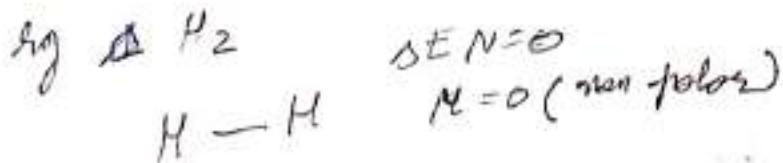
$\Delta E N = 0 \Rightarrow$ non-polar bond

$\Delta E N \neq 0 \Rightarrow$ polar

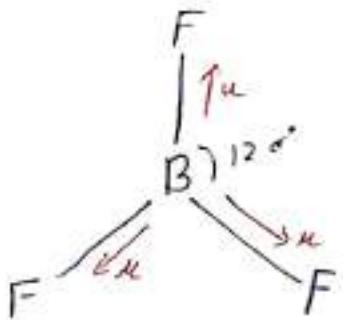
b) Polarity of molecules

$\Delta E N M = 0 \Rightarrow$ non-polar molecule

$\neq 0 \Rightarrow$ polar molecule

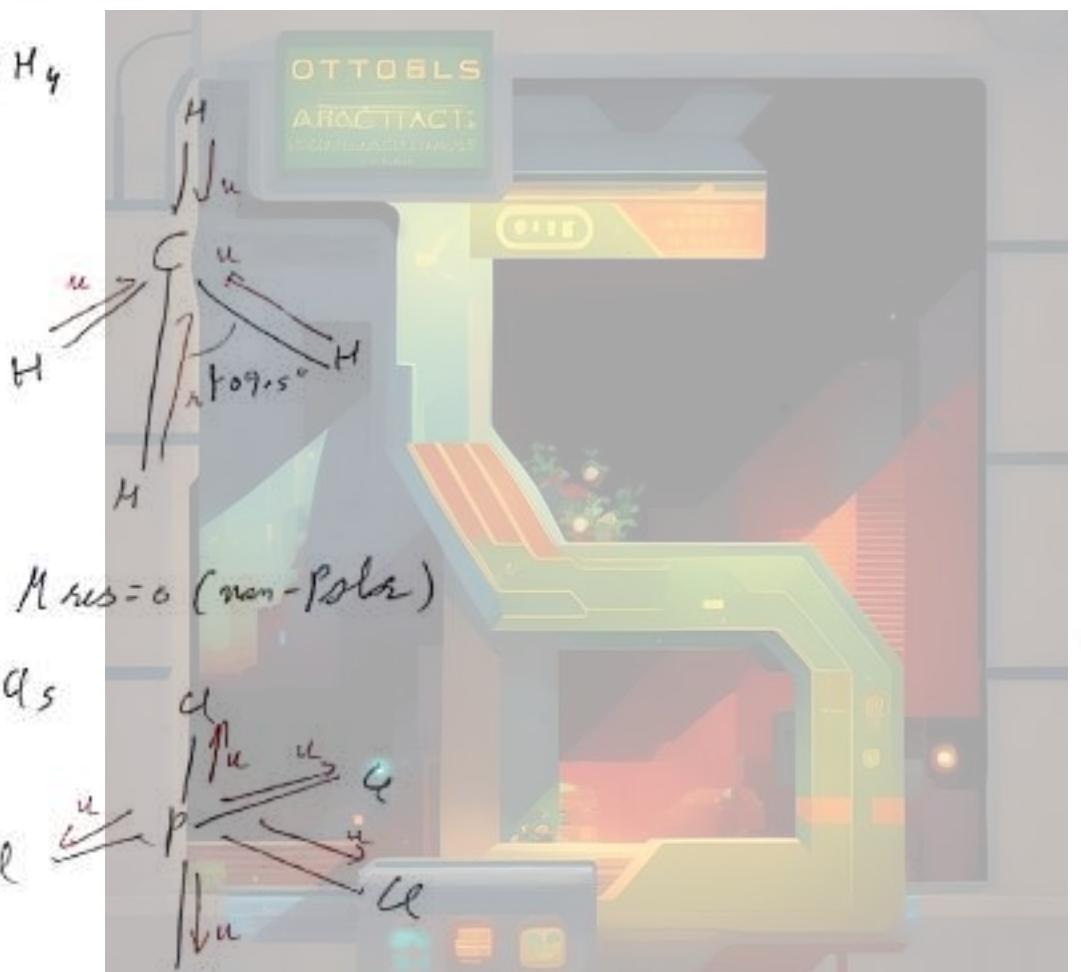


eg 3. BF_3

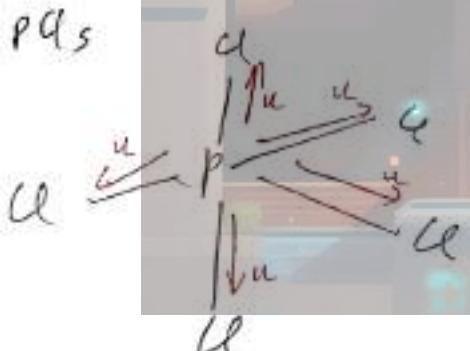


$M_{\text{res}} = 0$ (vector addition)
(non-Polar)

eg 4. CH_4

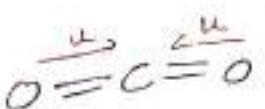


eg 5. PQ_5



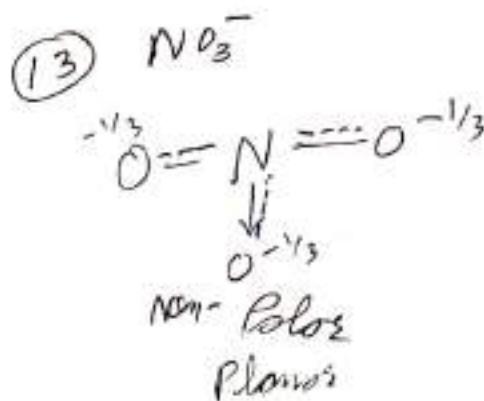
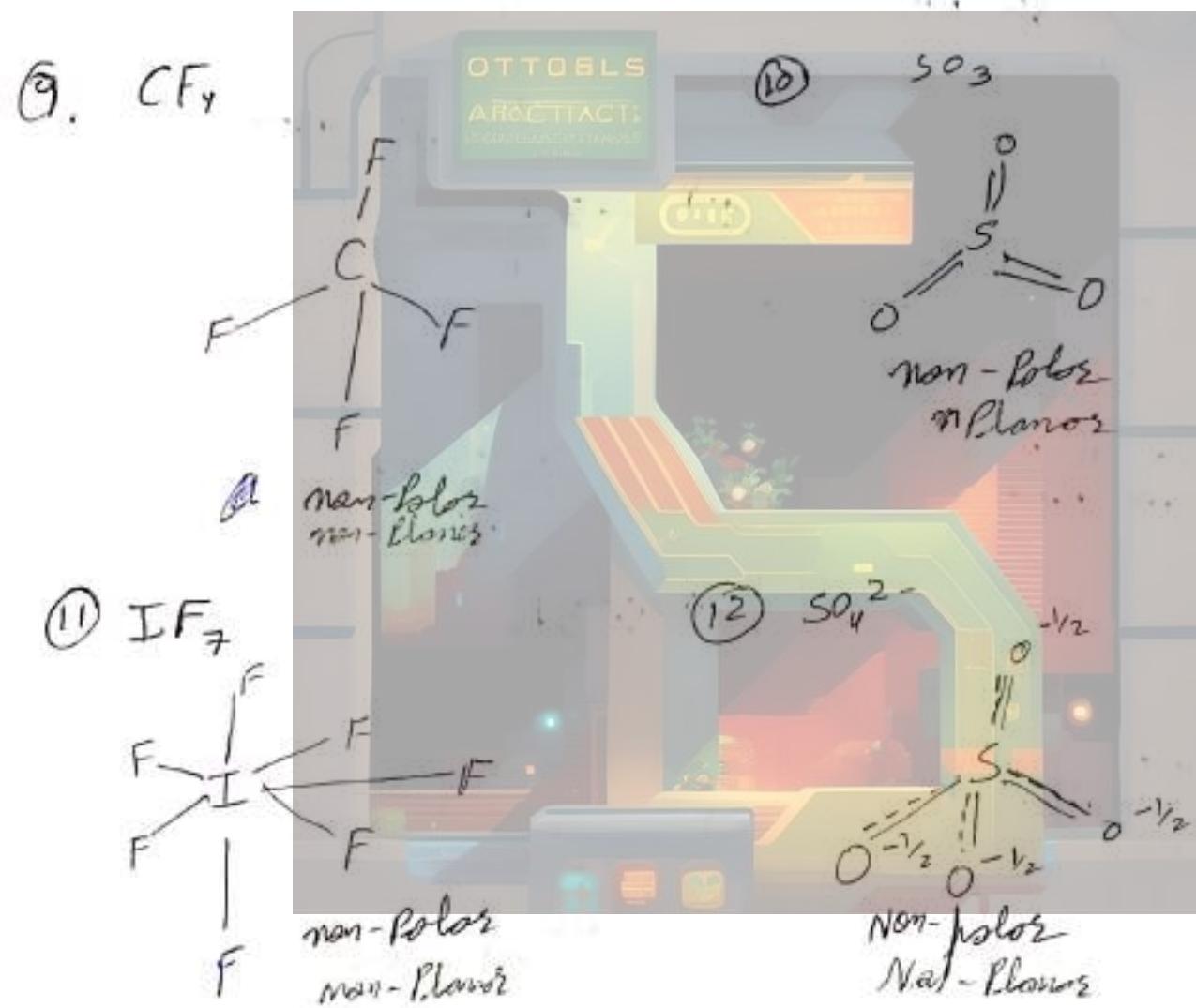
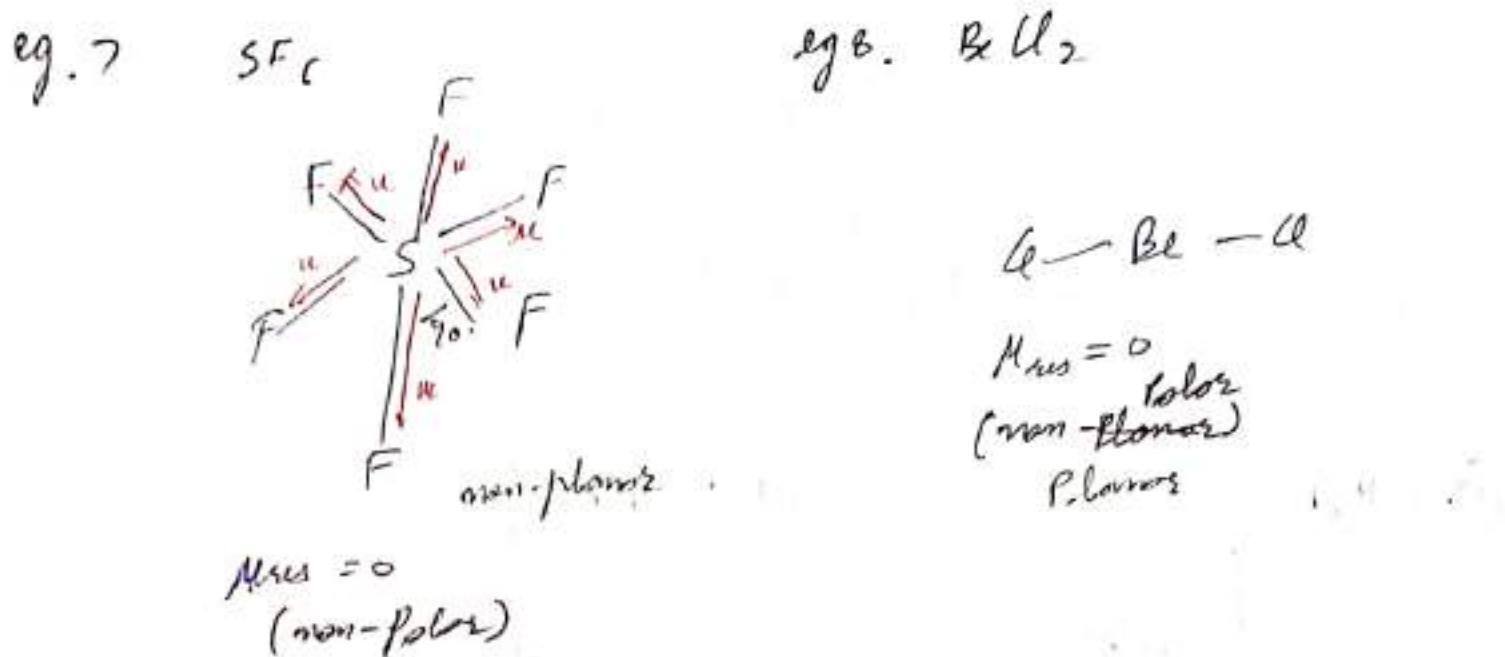
$M_{\text{res}} = 0$ (non-Polar)

eg 6. CO_2



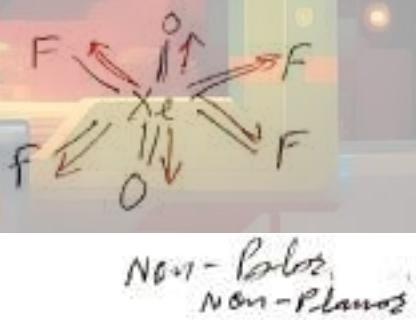
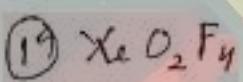
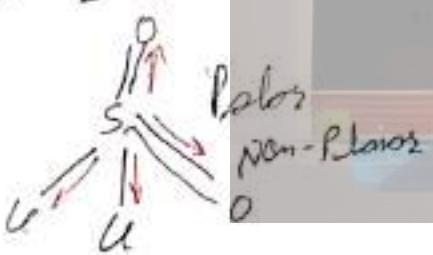
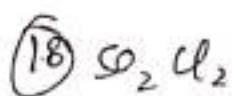
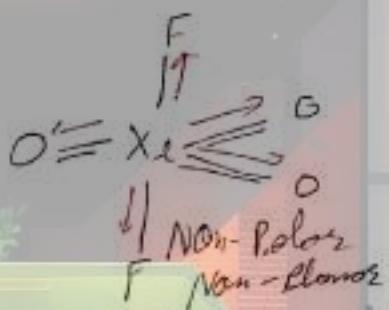
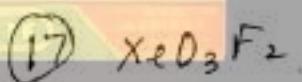
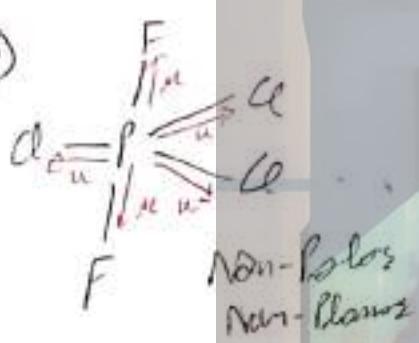
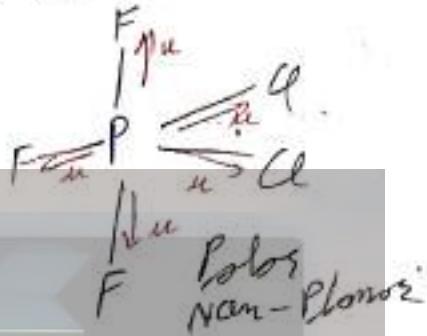
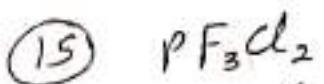
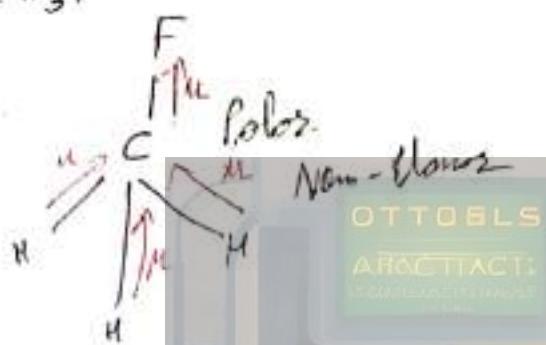
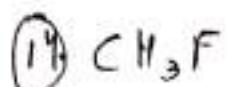
$M_{\text{res}} = 0$ (non-Polar)

linear



Note :-

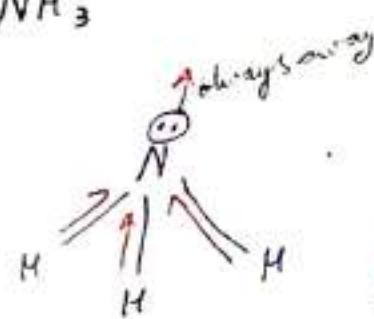
1. All regular geometries with no. L.P on central atom & having some surrounding atom are non-Polar.
2. All resonance Hybrid are non-Polar



NOTE:-

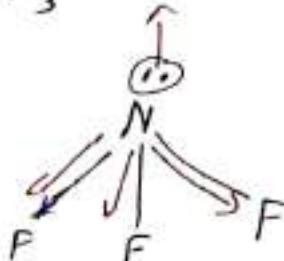
1. Dipole moment for a L.P is in the direction of orientation of L.P.
2. Dipole moment of a L.P can only be cancelled by other L.P dipole moment for this combination

(20) NH_3



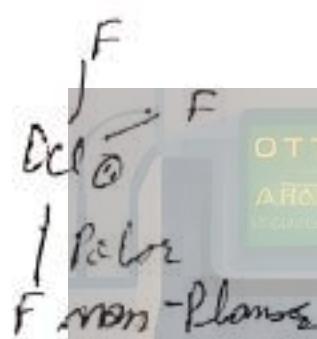
Polar
non-Planar

(21) NF_3

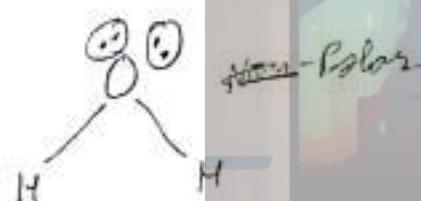


~~Never Polar~~
Non-Planar

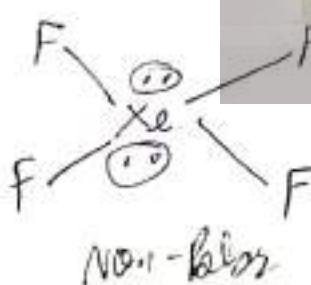
(22) ClF_3



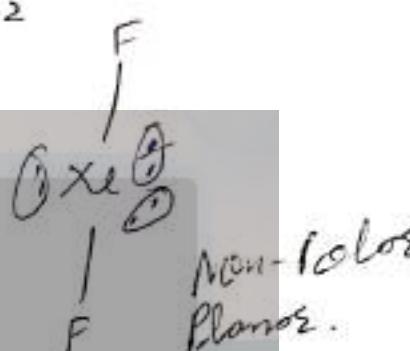
(24) H_2O



(26) XeF_4

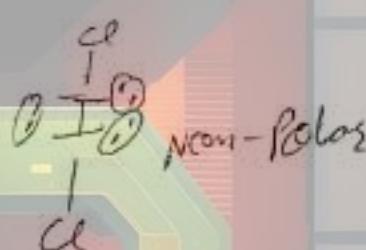


(23) XeF_2



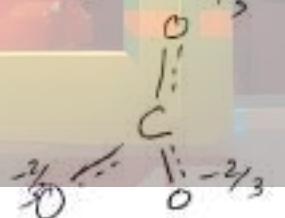
Non-Polar
Planar

(25) I_2Cl^-



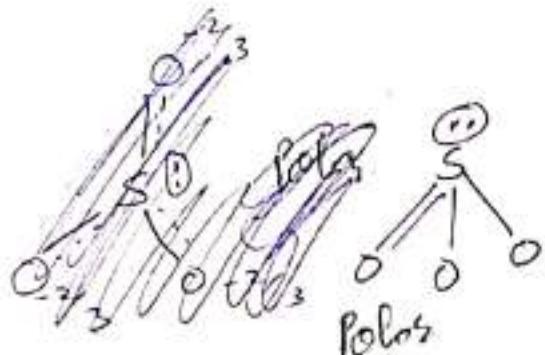
Cl

(27) CO_3^{2-}

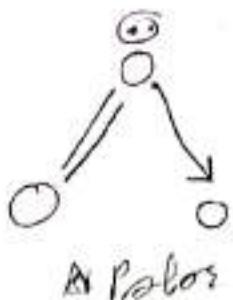


Non-Polar

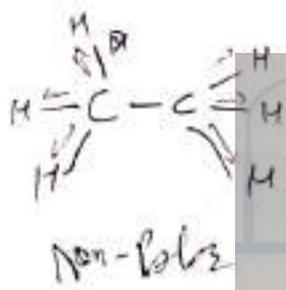
28 SO_3^{2-}



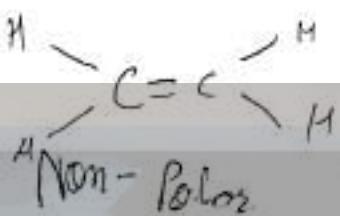
29 O_2



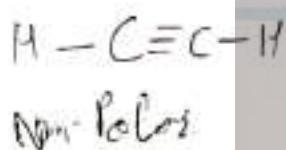
30 C_2H_6



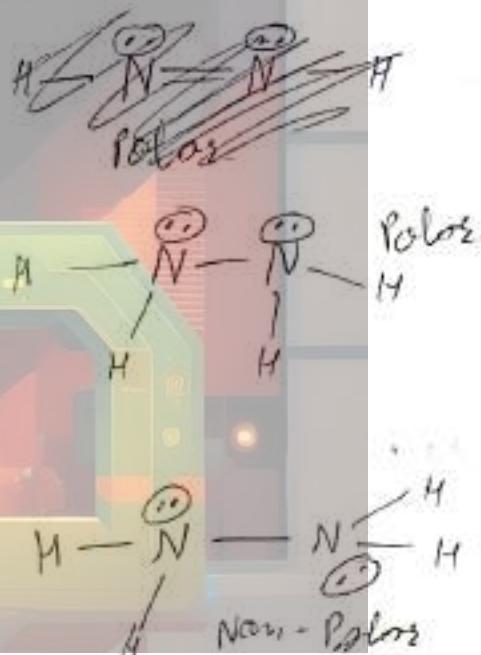
31 C_2H_4



32 C_2H_2



33 N_2H_4

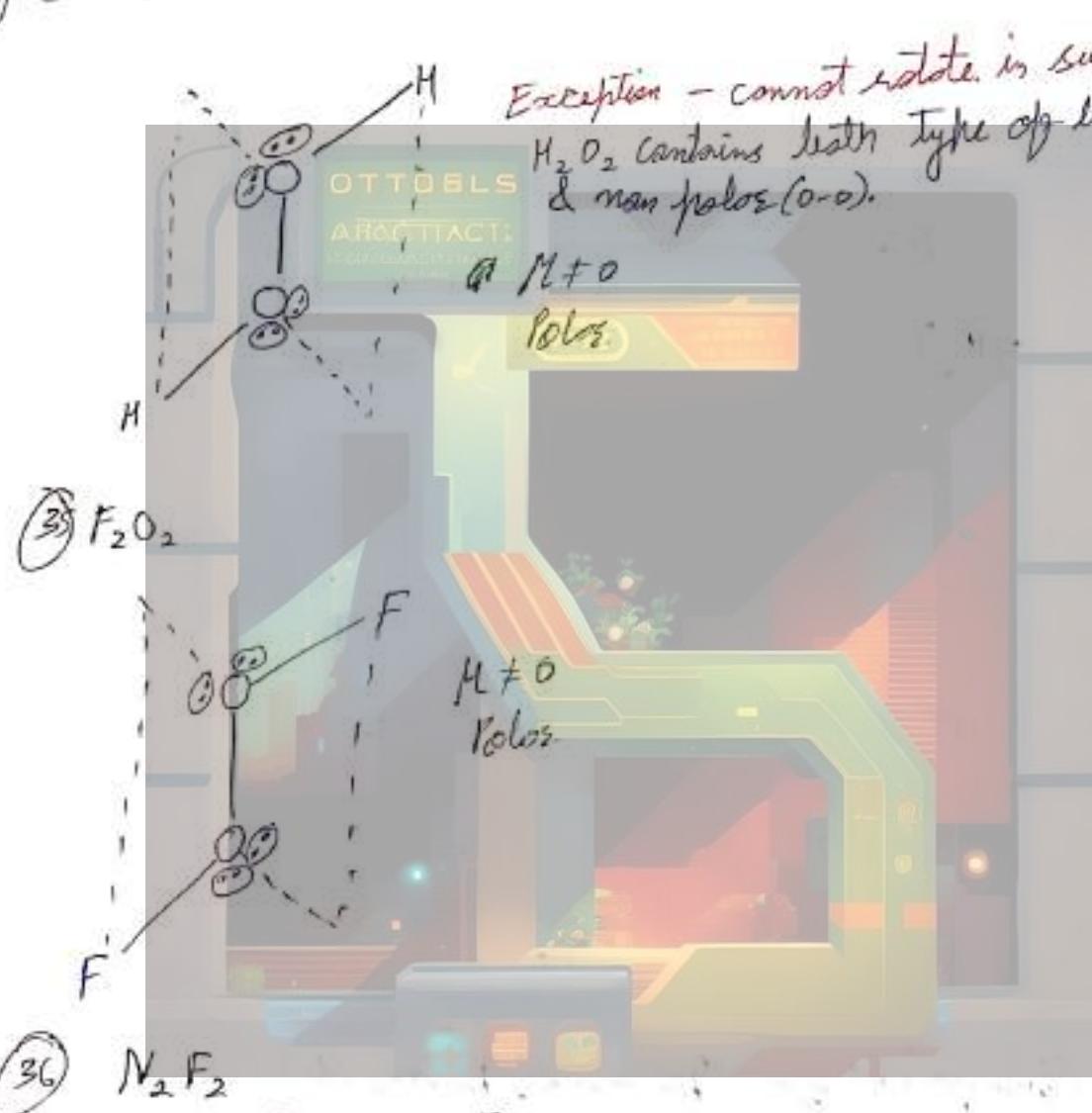


$\text{non-pole} + \text{pole} = \text{non-pole}$
Polar

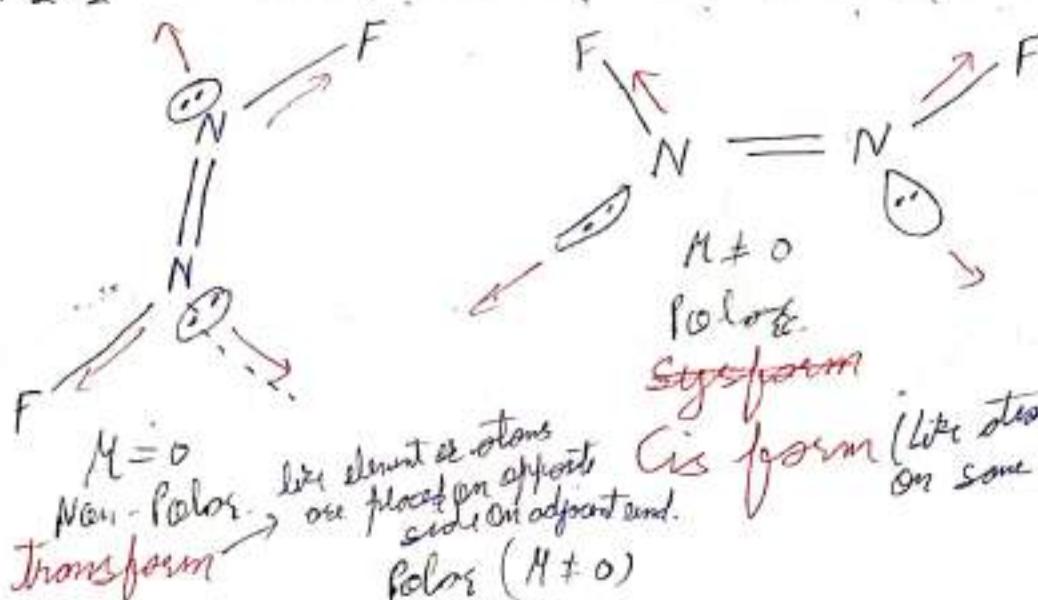
Note →

1. Single Bond rotates their position hence Position of lone pair is not fixed.
2. Rotation around single bond is possible but in multiple bond rotation is restricted.

e.g. (34) H_2O_2



(35) F_2O_2

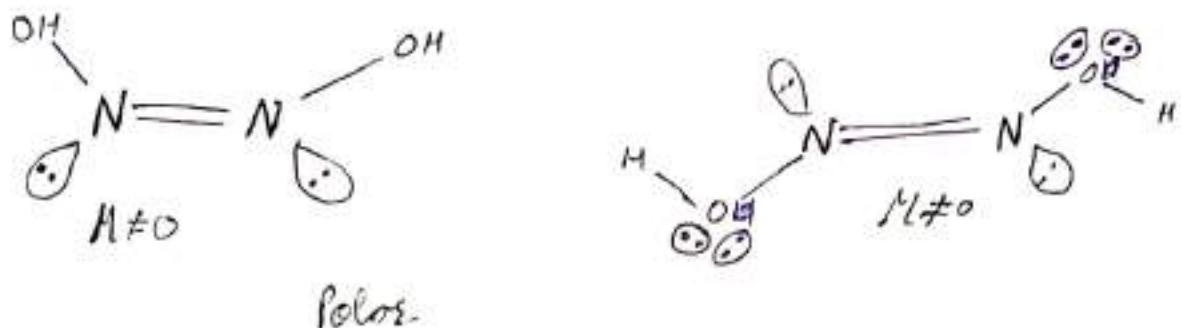


Transform

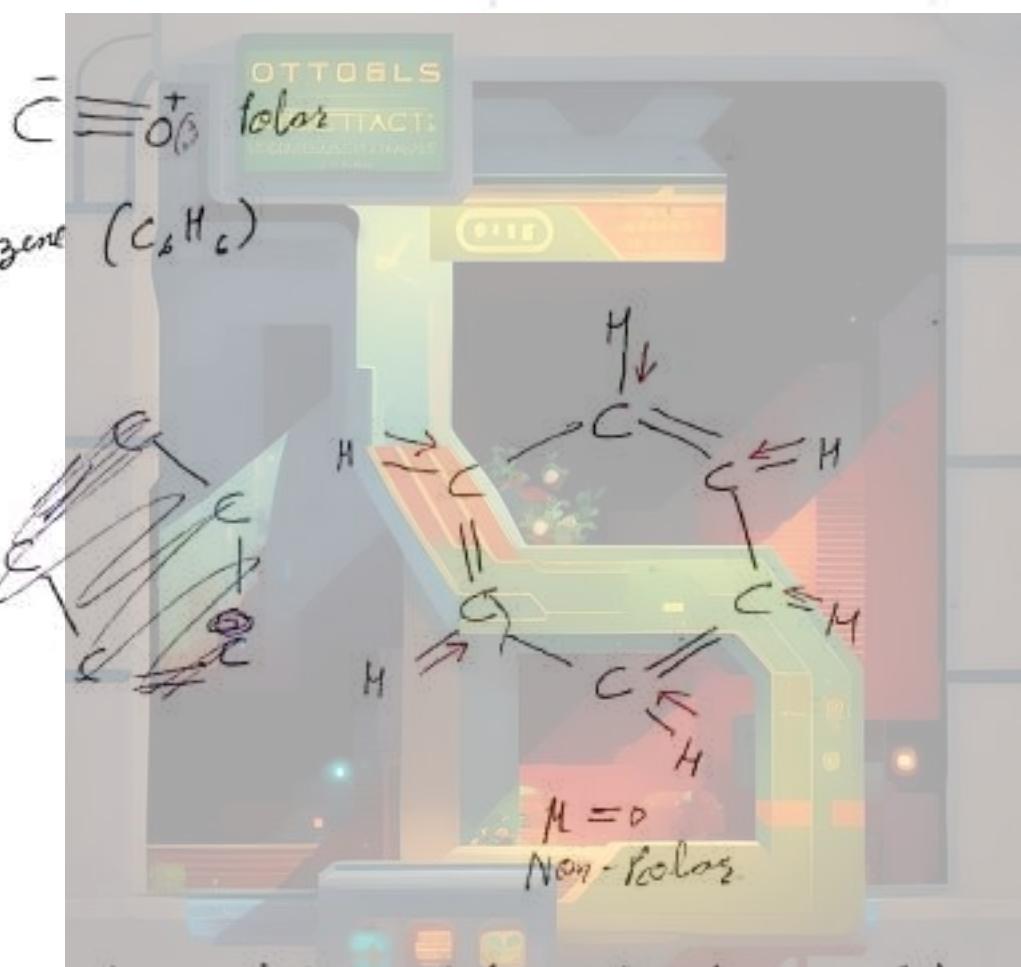
Polar (M+O)

(17)

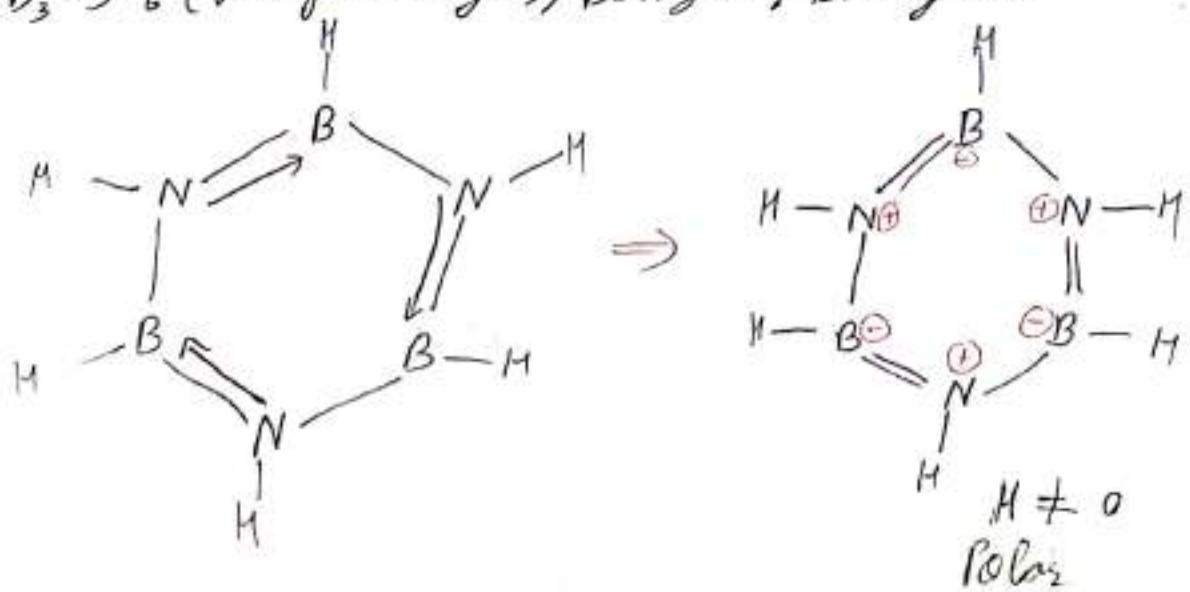
③7) $\text{H}_2\text{N}_2\text{O}_2$ or $\text{N}_2(\text{OH})_2$



③8) CO

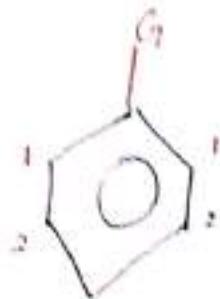


⑨) $\text{B}_3\text{N}_3\text{H}_6$ (Inorganic Benzene)/Borazine / Borazode)



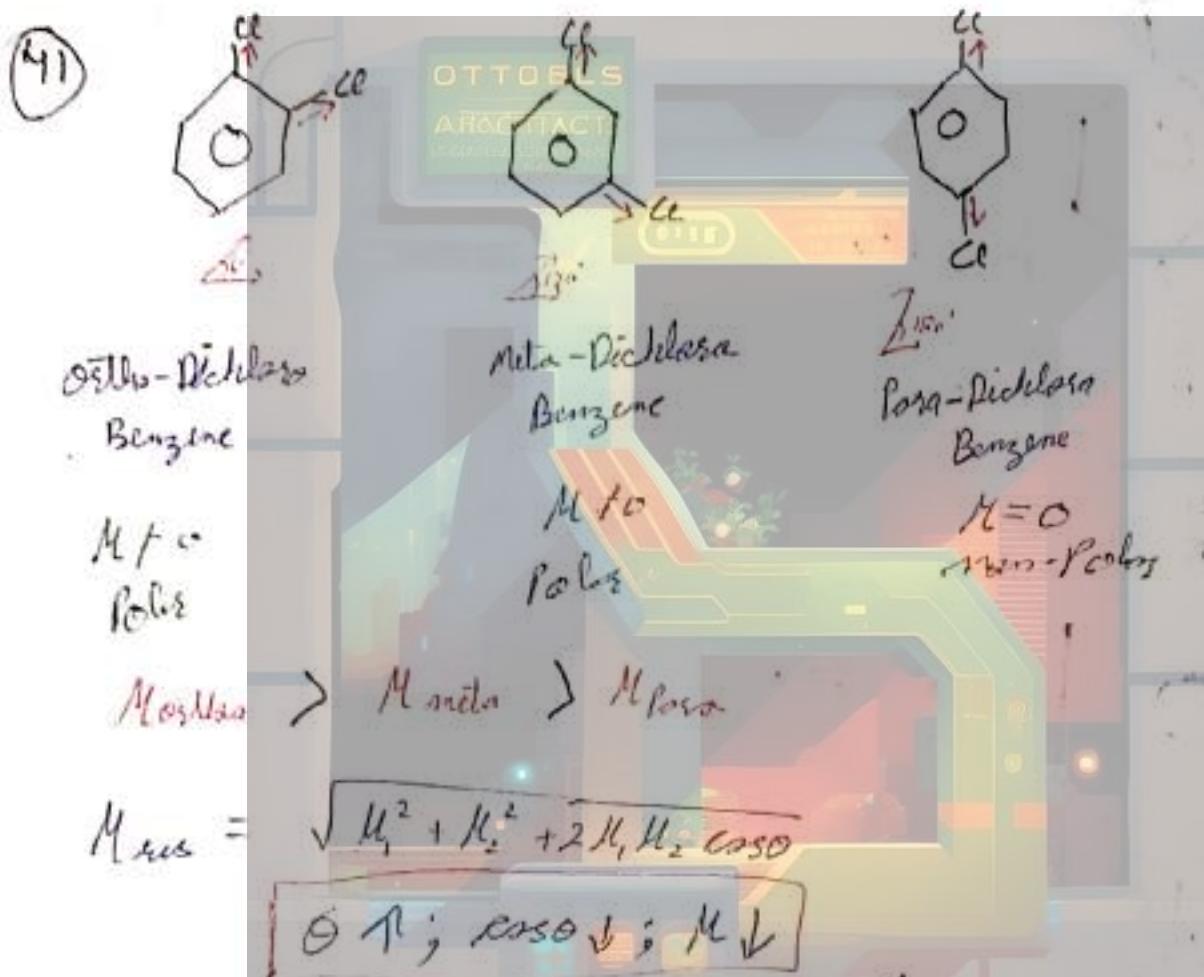
Note :-

1. Organic Benzene & Inorganic benzene both are iso-structural & iso-electronic.

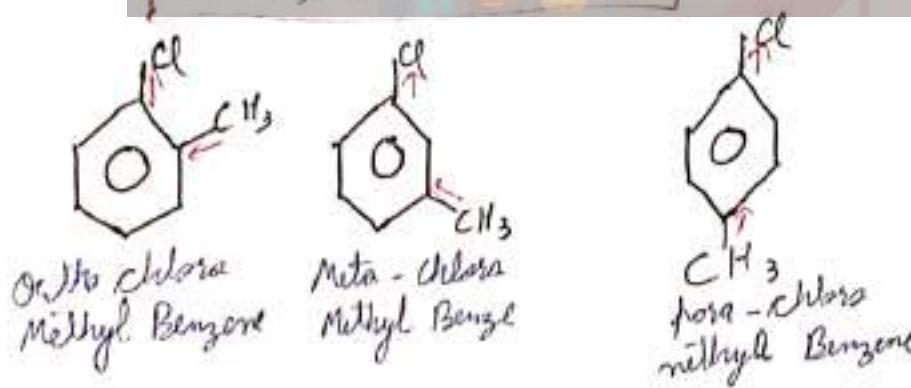


1 → either position
2 → Meta position
3 → Para position

with respect to G1



(42)



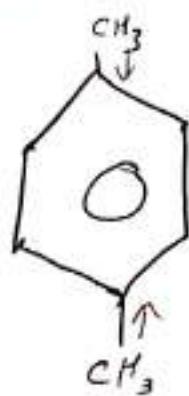
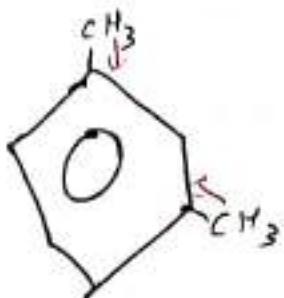
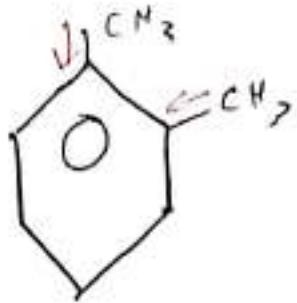
$\mu \neq 0$

$\mu \neq 0$

$\mu \neq 0$

Mortho $\mu < \mu_{\text{meta}}$ $\mu < \mu_{\text{para}}$
polar all

(43)



μ_{dipole} > μ_{meta} > μ_{para}

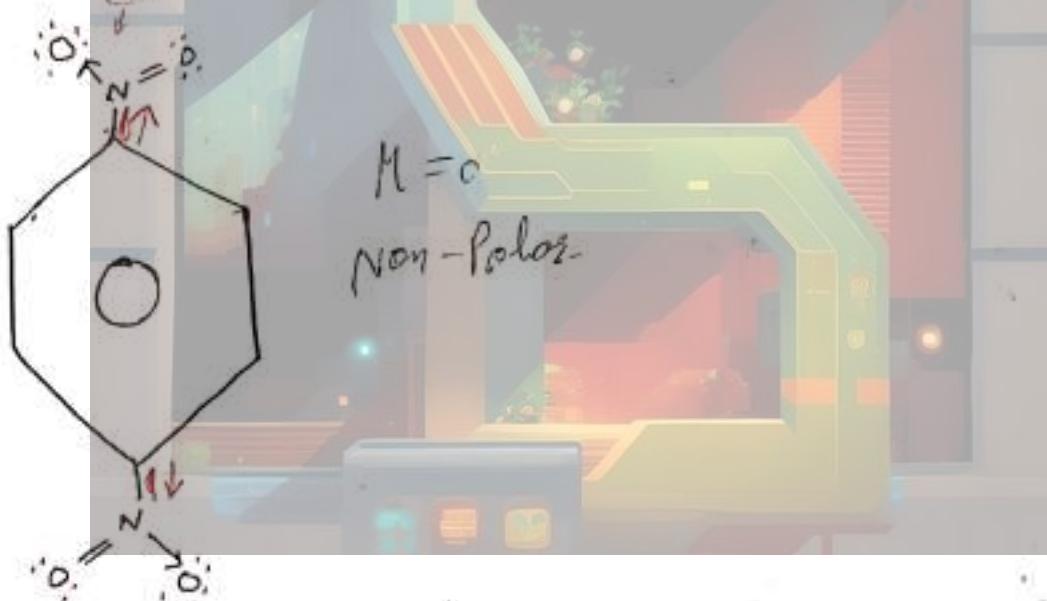
$\mu \neq 0$ Polar

$\mu = 0$ Non-Polar

(44)



(45)



(120)

Q1 Identify Polar & Non-Polar (H.W)

- | | | | |
|----------------------------|----------------------|---------------------------|----------------------------|
| ① PCl_3F_2 | ⑥ NO_3^- | ⑩ Cl_2O_2 | ⑯ SO_2Cl_2 |
| ② XeO_3F_2 | ⑦ CO_3^{2-} | ⑫ BeF_2 | ⑰ IO_2^- |
| ③ XeO_3 | ⑧ SO_2 | ⑬ XeO_4 | |
| ⑨ XeF_5^- | ⑨ SO_3 | ⑭ BFCl_2 | |
| ⑤ XeF_4 | ⑩ SOCl_2 | ⑮ N_3H | |

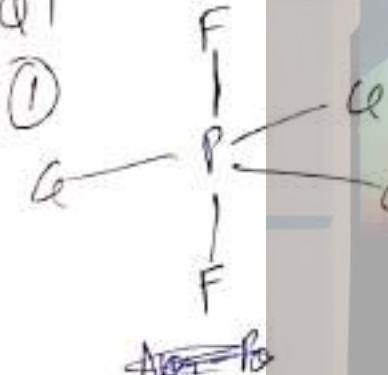
polar species.

Q2. No. of non polar & non-polar species.

- ① ClO_3^-
- ② SO_2F_2
- ③ SO_4^{2-}
- ④ CH_2F_2
- ⑤ PO_4^{3-}
- ⑥ IO_4^-

- | | |
|---|---------------------------|
| ⑦ Colinear $\text{Cl}_2\text{O}_6(s)$ | ⑫ CO |
| ⑧ Linear $\text{N}_2\text{O}_5(s)$ | ⑬ NO_2^- |
| ⑨ XeFs ABSTRACTS | ⑭ IF_4^{\ominus} |
| ⑩ I_3^{\oplus} | ⑮ IF_7 |
| ⑪ I_3^{\ominus} | ⑯ SF_6 |

Q1



Non-Polar ✓
Non-Planar ✓

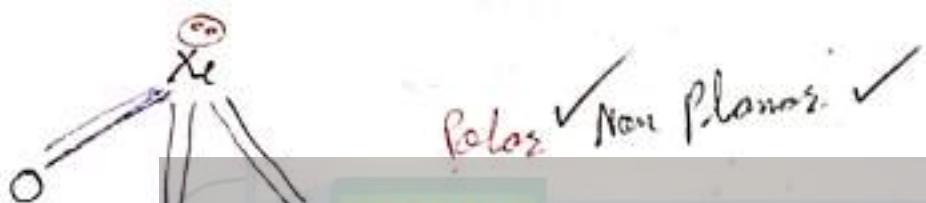
②



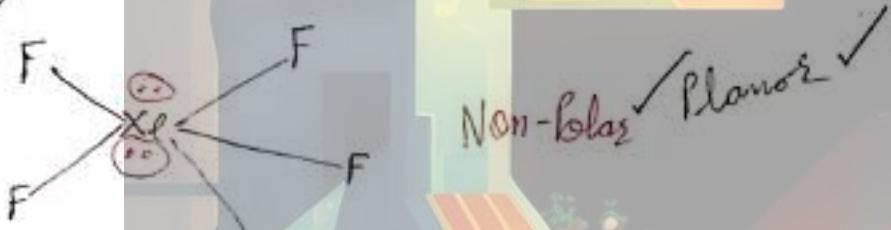
(2)



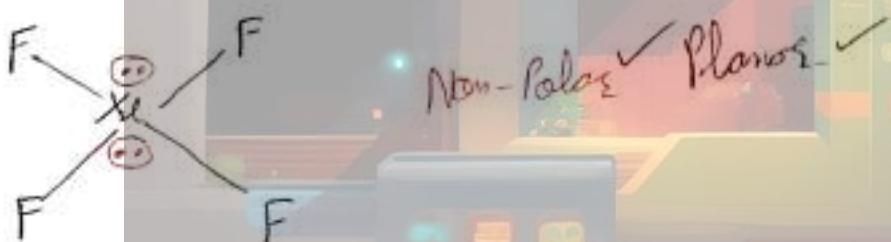
(3)



(4)

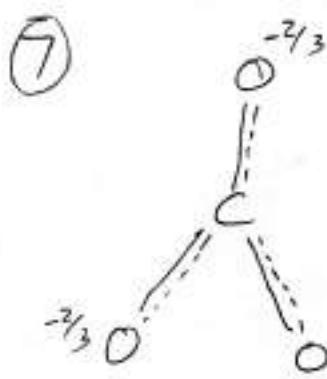


(5)

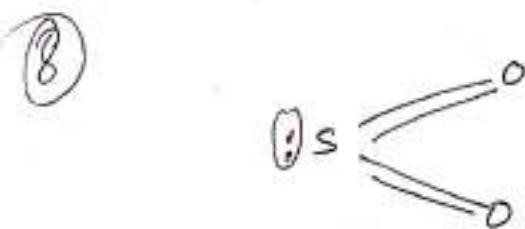


(6)

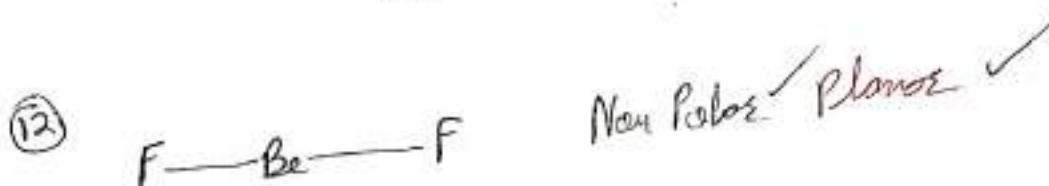
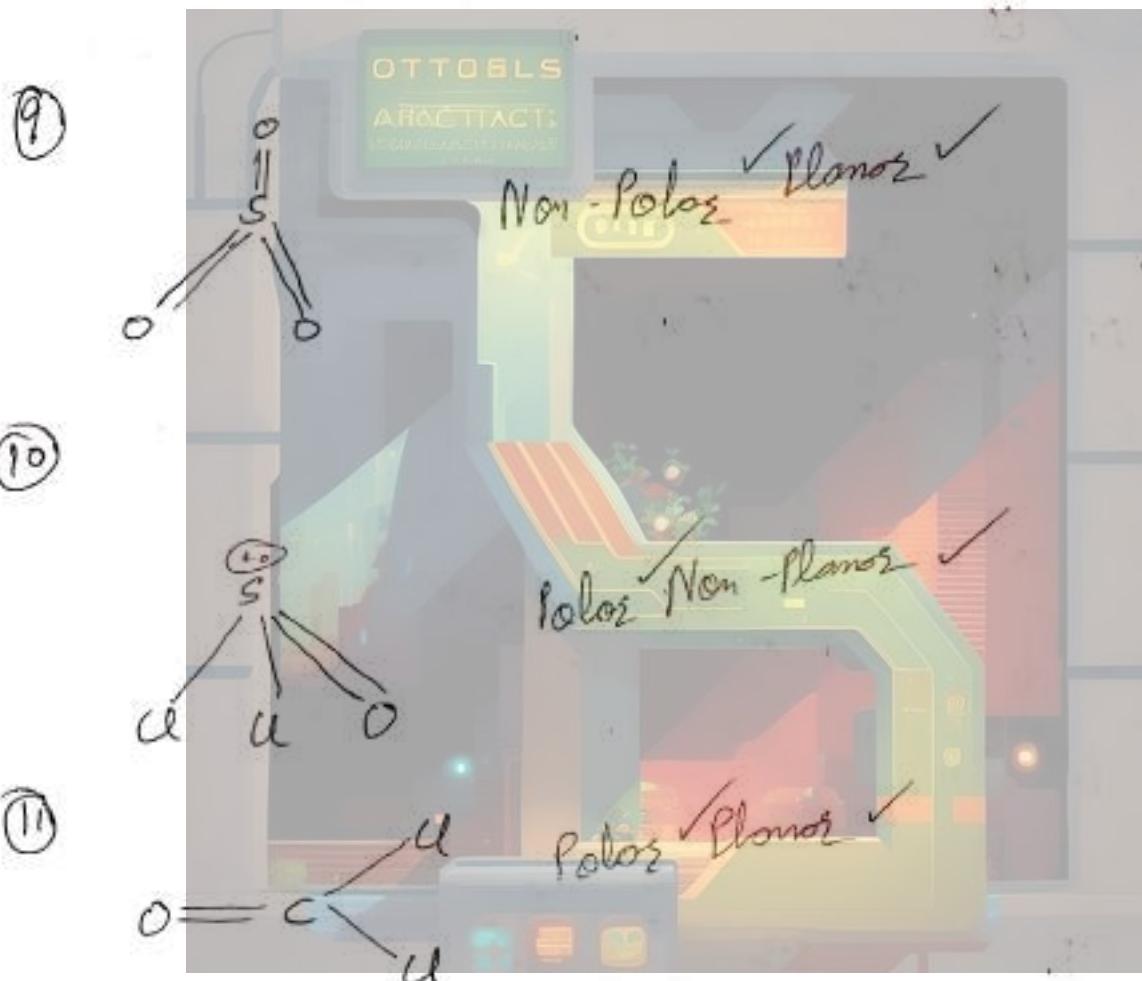




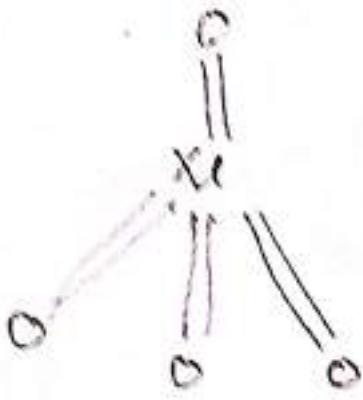
Non-Polar ✓ Planar ✓



Polar ✓ Planar ✓



(13)



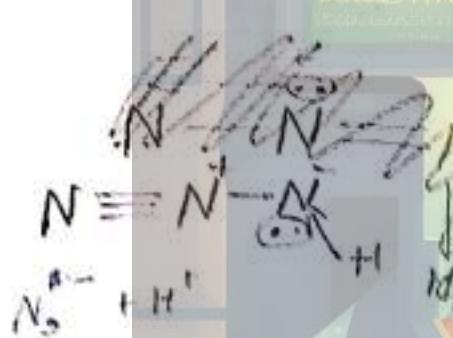
Non
Non-polar ✓ Non-polar ✓

(14)



Polar ✓ Planar ✓

(15)



Polar ✓ Dimer ✓

(16)



Polar ✓ Non-polar ✓

(17)



Polar ✓ Planar ✓

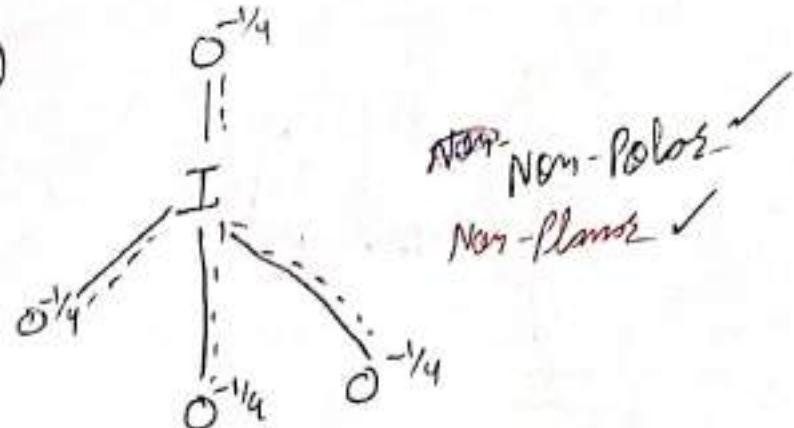
(124)

Q2.

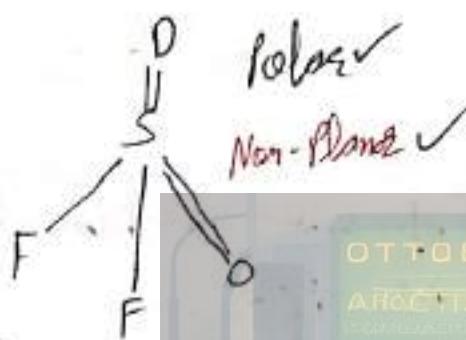
①



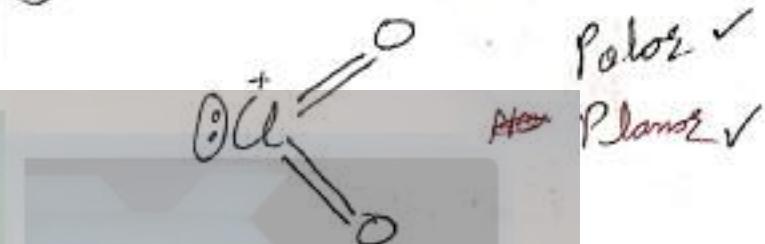
⑥



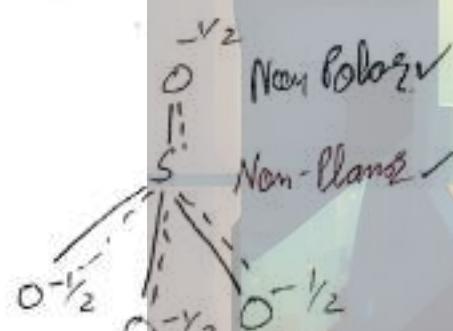
②



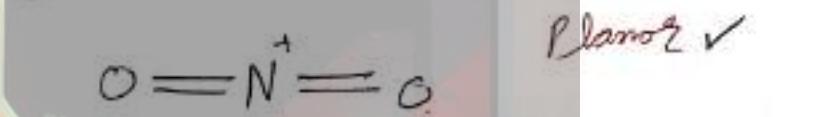
⑦



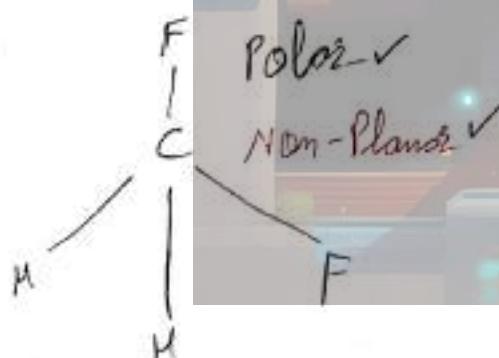
③



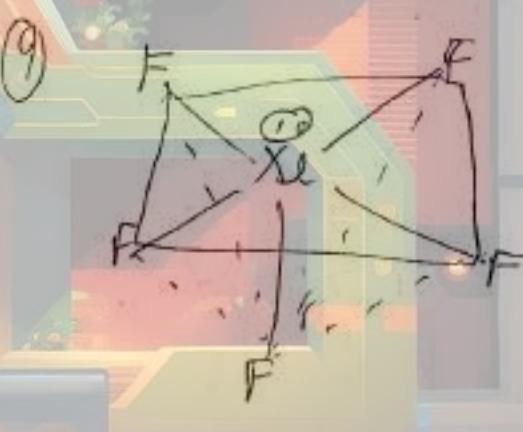
⑧



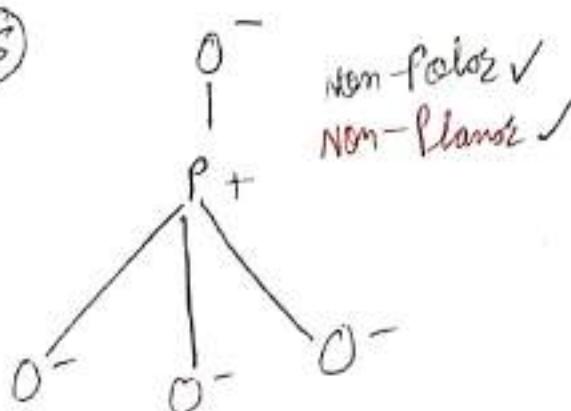
④



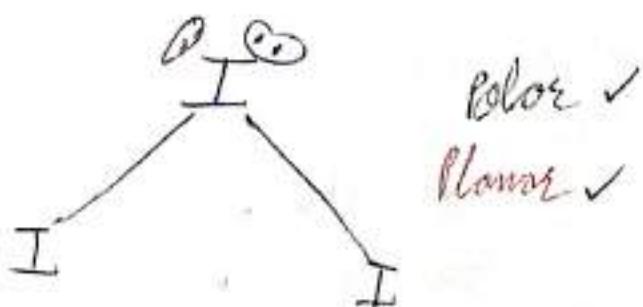
⑨



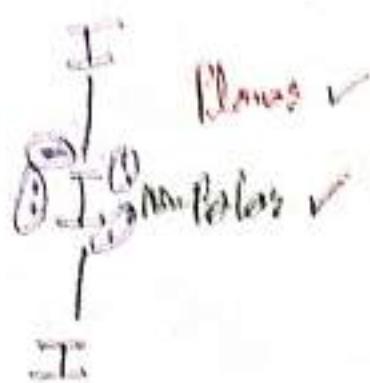
⑤



⑩



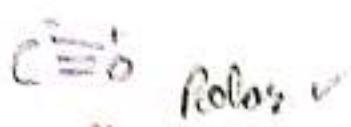
(11)



(16)



(12)



Planes ✓
H

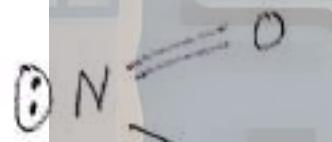
Non-Polar ✓

(13)



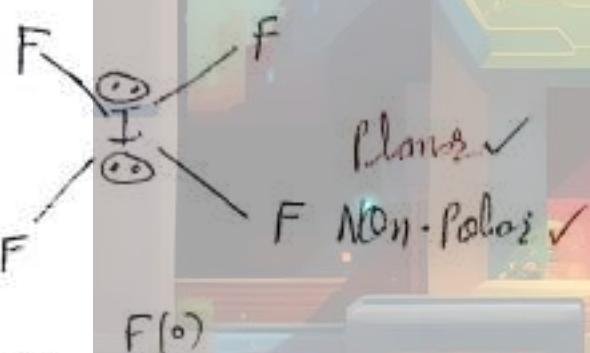
OTTOBL'S
FACTS

5 species are non-Polar b
Non-Polar

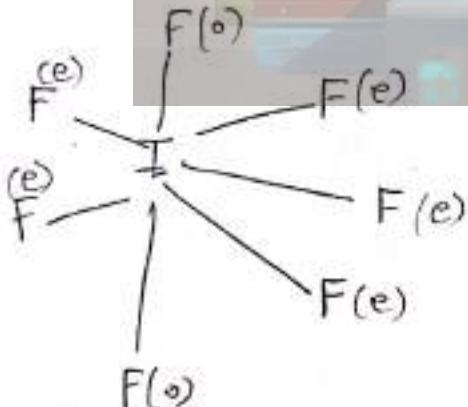


Planes ✓
Non-Polar ✓

(14)



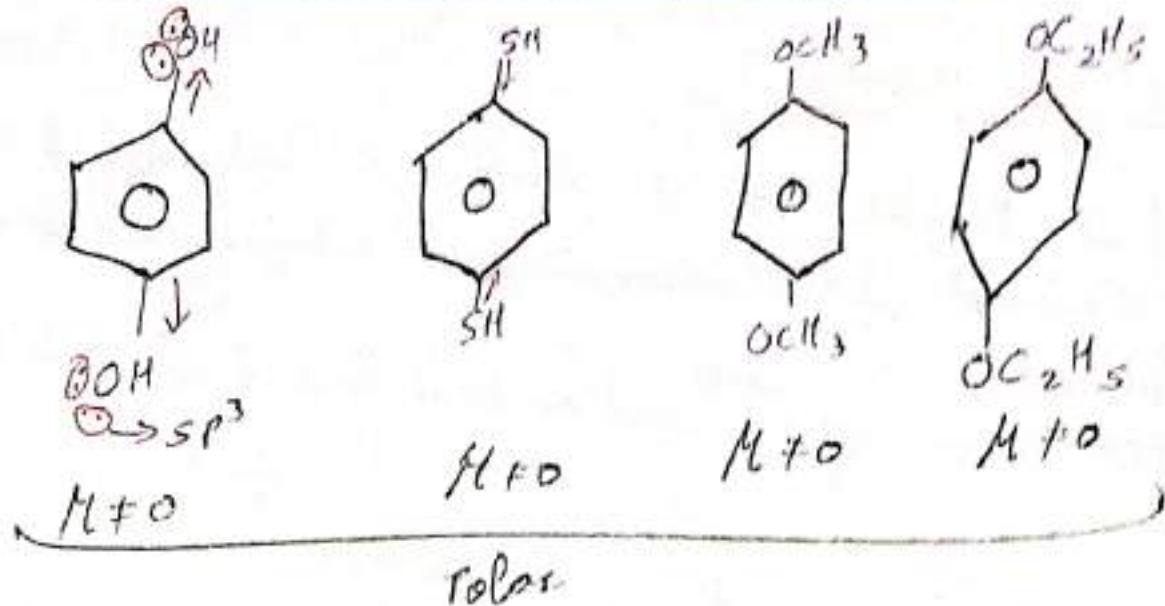
(15)



Non-Polar ✓
Non-Planar ✓

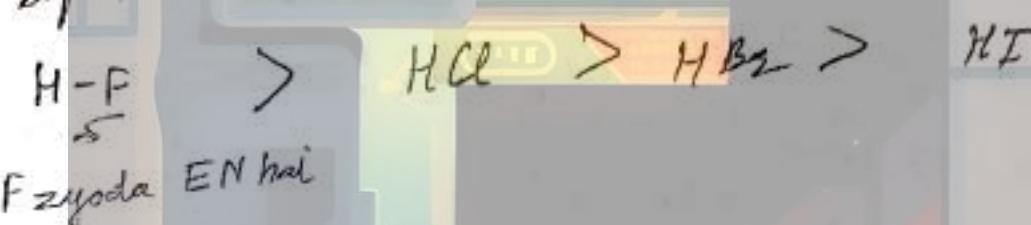
(126)

(46)

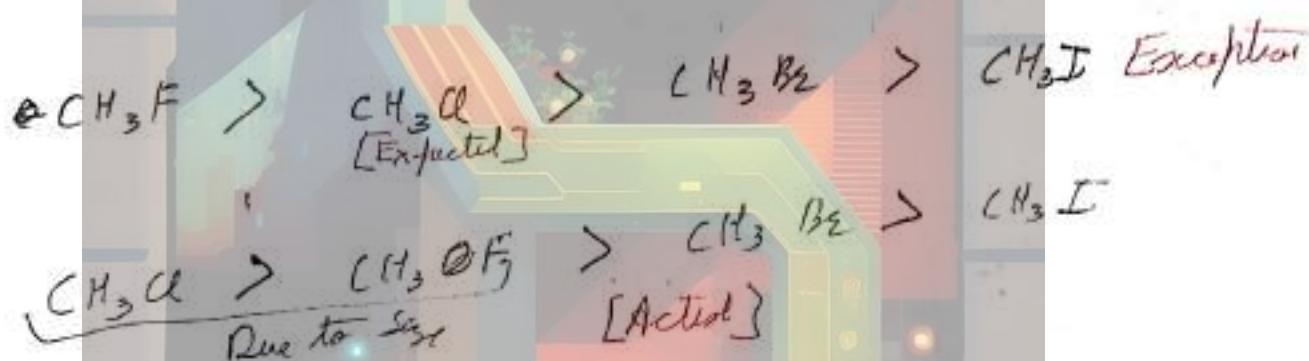


(47)

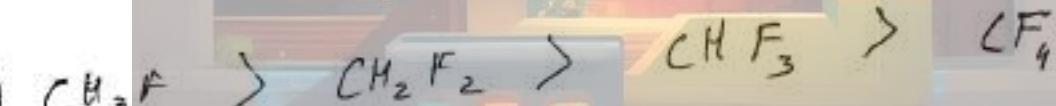
dipole moment order



(48)



(49)



H.W. 20-6-24

O-1 [18-30]

O-2 [12-18]

S-1 [1-9, 12]

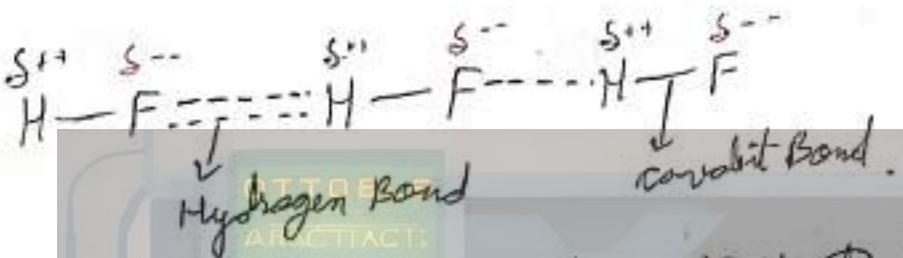
S-2 [1, 2, 5, 6, 7, 8, 9, 10]

(127)

Hydrogen Bonding

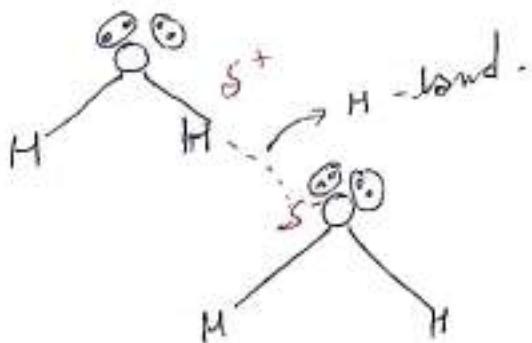
- It is the weak force of attraction between hydrogen & more EN element like Fluorine, Oxygen, Nitrogen & sometimes Chlorine.
- The amount of ^{Energy} released on bond formation is 8 to 42 kJ/mol

e.g.

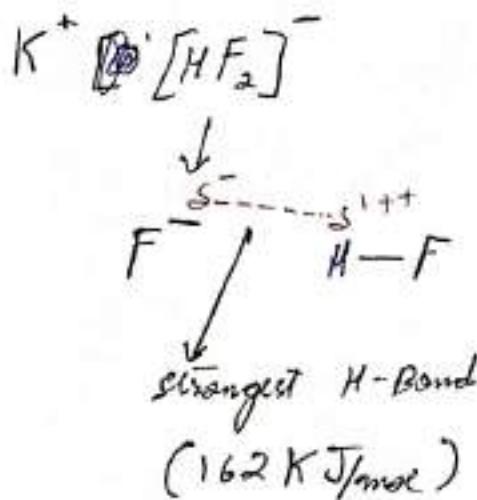


→ If Hydrogen is attached to strong EN atom (F, N, O, sometimes Cl) comes close to another strong EN atom. (F, N, O, Cl) Then interaction developed between them is called Hydrogen Bond.

e.g.



Strongest H-Bond $KHF_2(s)$



Strength of Hydrogen Bonding

- ① H.B.S \propto Charge accumulation on atom [E-N] on Hydrogen.
- ② If charge accumulated on Hydrogen atom is some then Hydrogen bond Strength is directly proportional to Electron danding Tendency of neighbouring atom.

H.B.S \propto EDL danding Tendency of neighbouring atom.

- Eg ① a) H-F \cdots S⁻ H-O
b) H-F \cdots S⁻ H-N
c) H-F \cdots S⁻ H-F

a > b > c

- d) H-O \cdots S⁻ H-F
e) H-O \cdots S⁻ H-N
f) H-O \cdots S⁻ H-N

~~b > c~~

d > e > f

- g) H-N \cdots S⁻ H-F
h) H-N \cdots S⁻ H-O
i) H-N \cdots S⁻ H-N

g > h > i

g > d > a > h > e > b > i > f > c

- Q a) $\text{H}-\ddot{\text{O}}-\cdots \text{H}-\text{F}$
 b) $\text{H}-\ddot{\text{N}}-\cdots \text{H}-\text{O}$
 c) $\text{H}-\ddot{\text{F}}-\cdots \text{H}-\text{N}$
 d) $\text{H}-\ddot{\text{F}}-\cdots \text{H}-\text{f}$

~~a>d>b~~

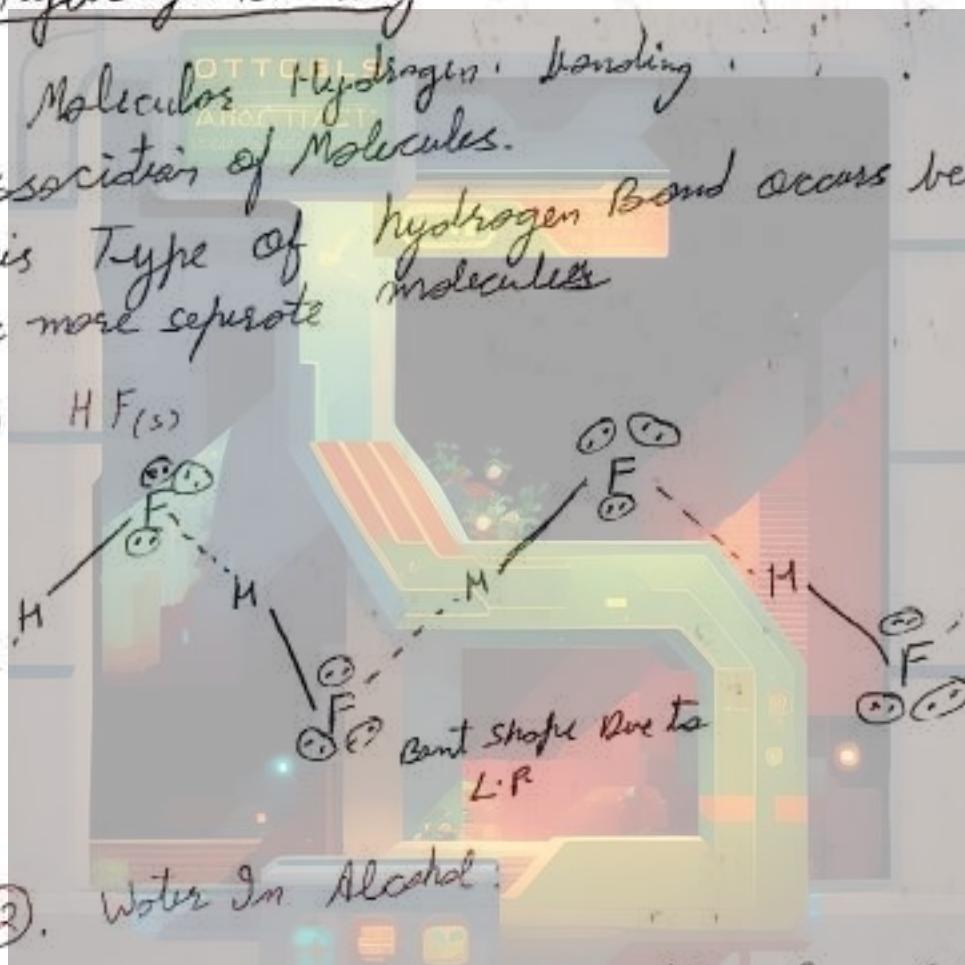
~~c>b>d~~

a>d>b>c

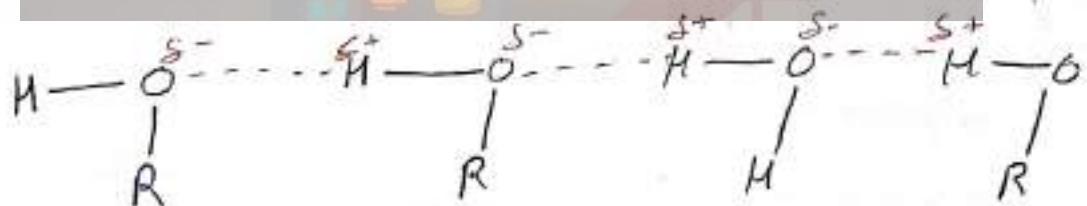
Types of Hydrogen Bonding

- ① Inter. Molecular Hydrogen Bonding
 → Association of Molecules.
 → This Type of hydrogen Bond occurs between two or more separate molecules

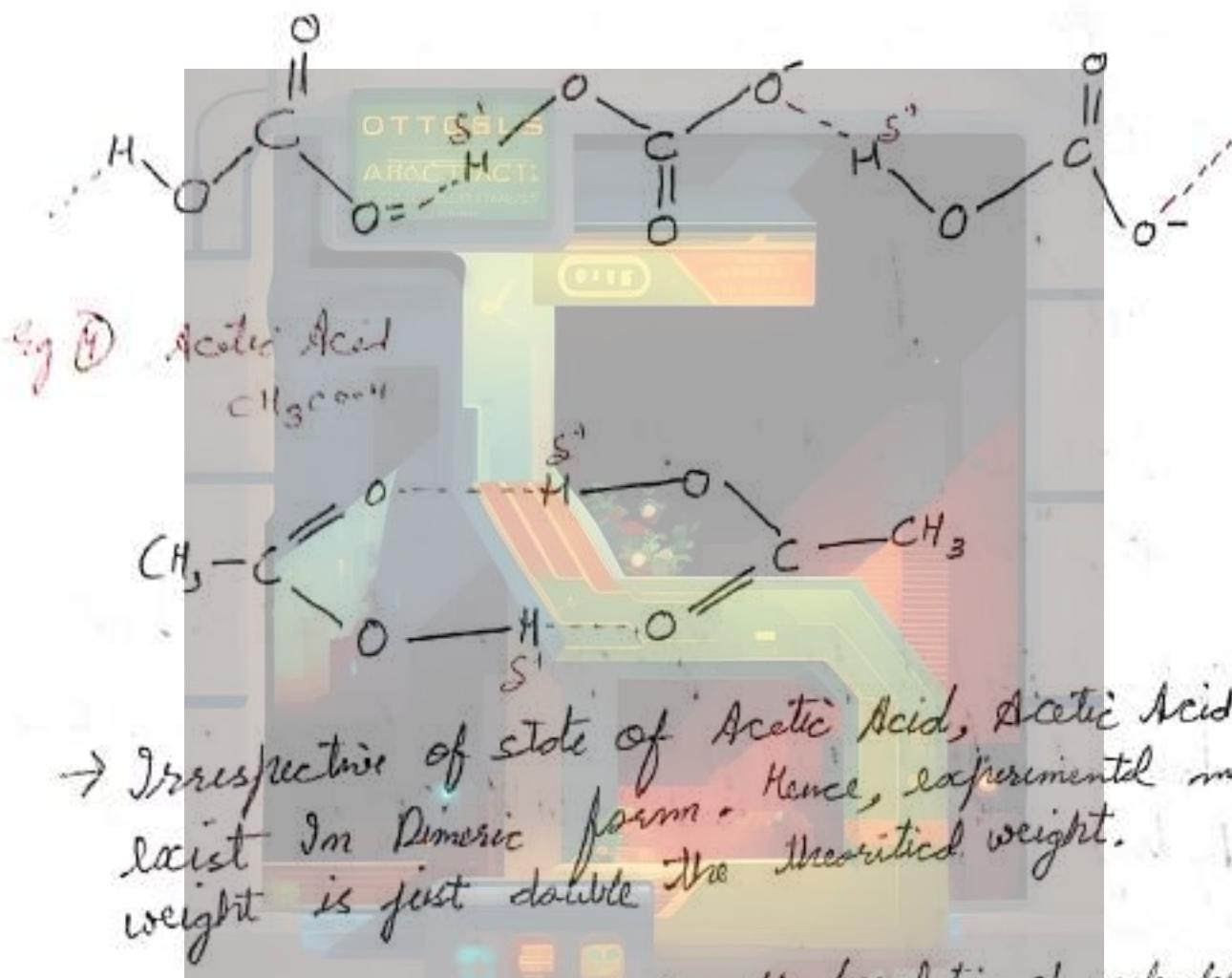
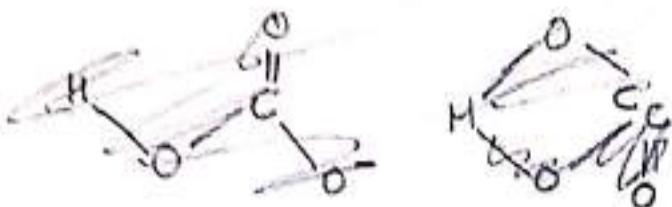
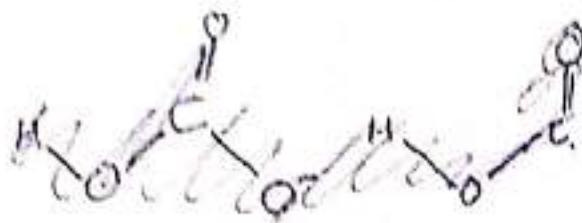
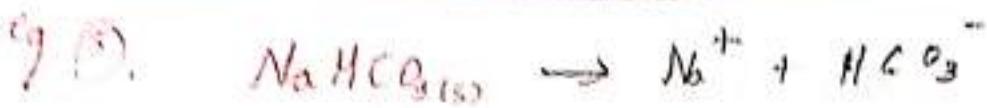
eg ①



eg. ②. Water In Alcohol

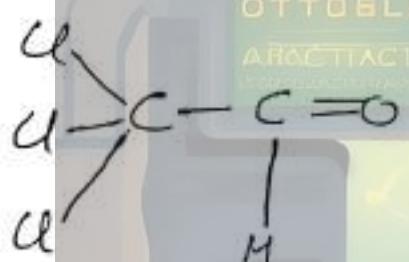
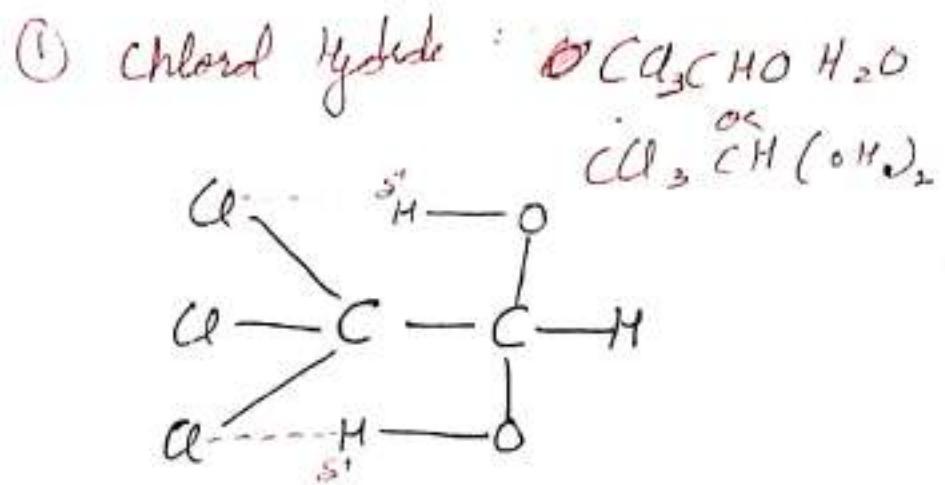


$\text{R} \rightarrow$ Alkyl group ($\text{CH}_3, \text{C}_2\text{H}_5$ etc)



(2) Intra Molecular Hydrogen Bonding / Chelation of molecule / Ring Formation.

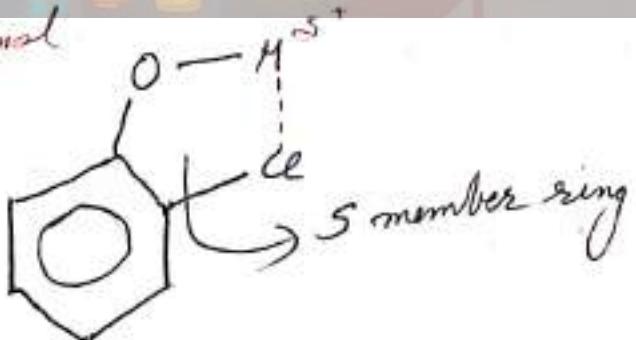
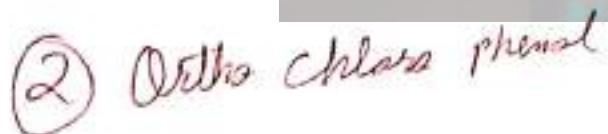
→ This type of hydrogen bonding occurs within a molecule.

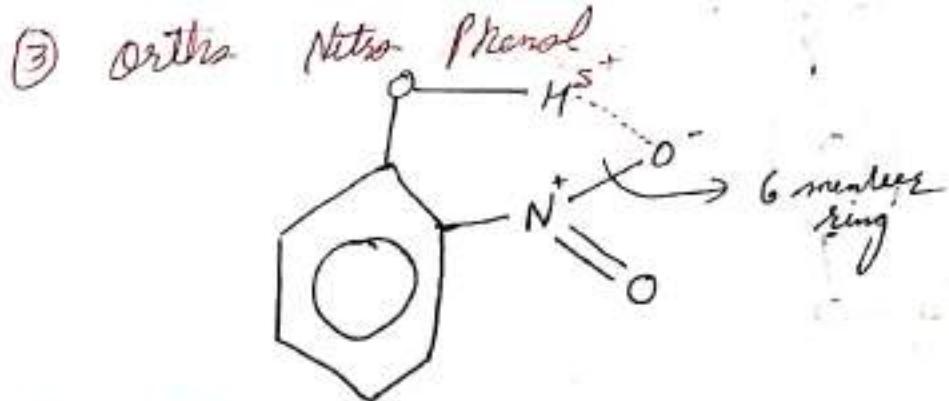


No hydrogen

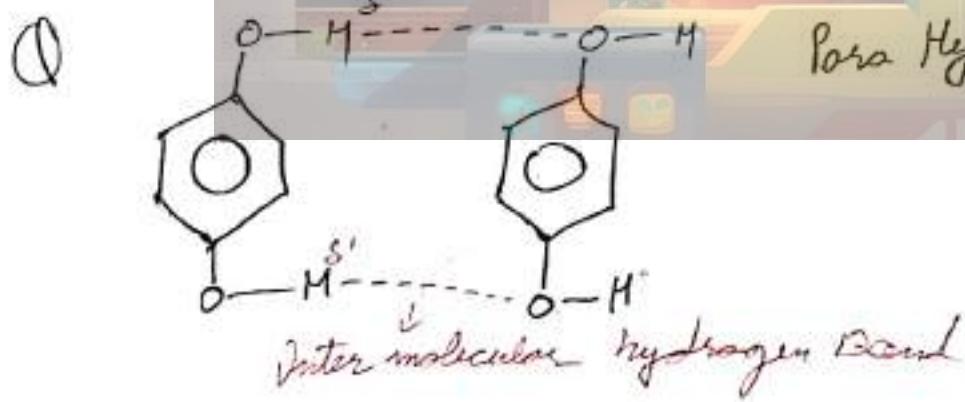
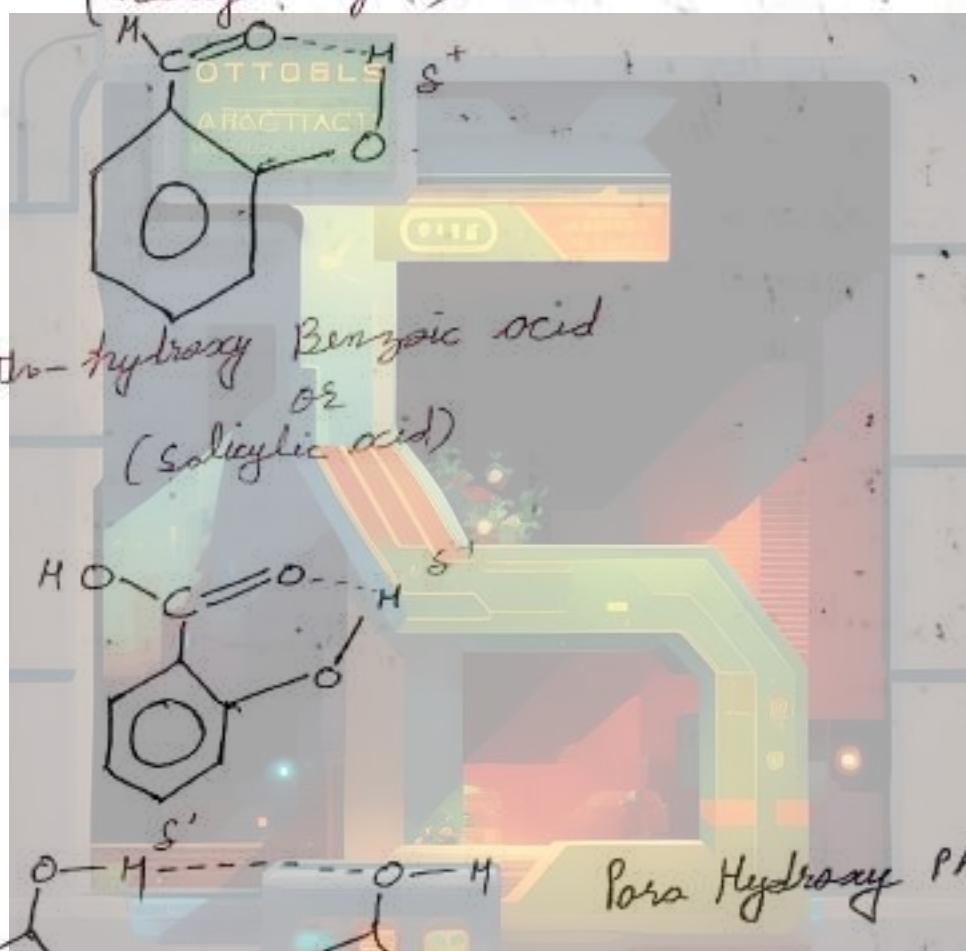
Bonding as H is not attached to
 $\text{C}, \text{N}, \text{O}$ or F)

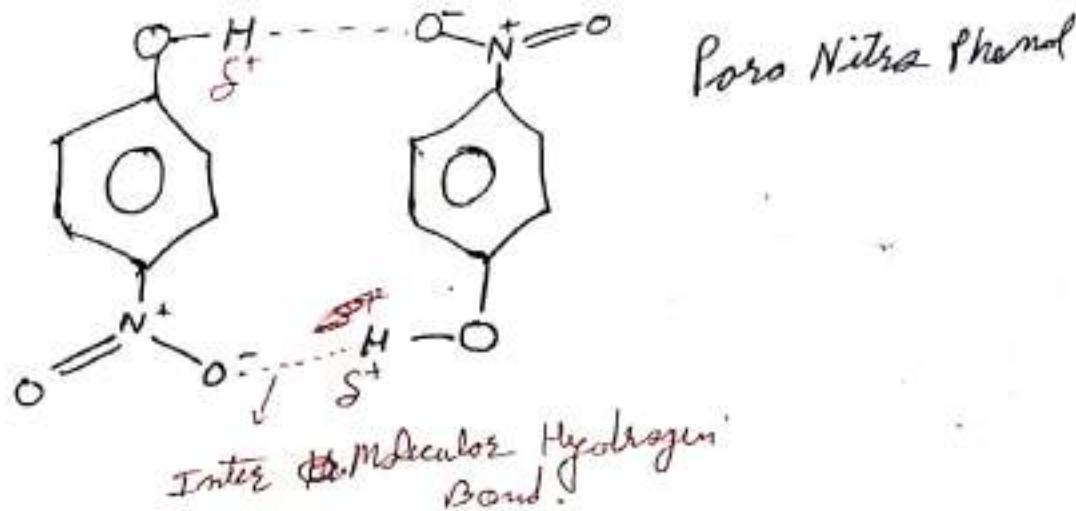
→ Chloral Hydrate forms Hydrogen Bond but Chloral does not form hydrogen bonding because in chloral hydrogen atom is attached with Sp^2 hybridised carbon





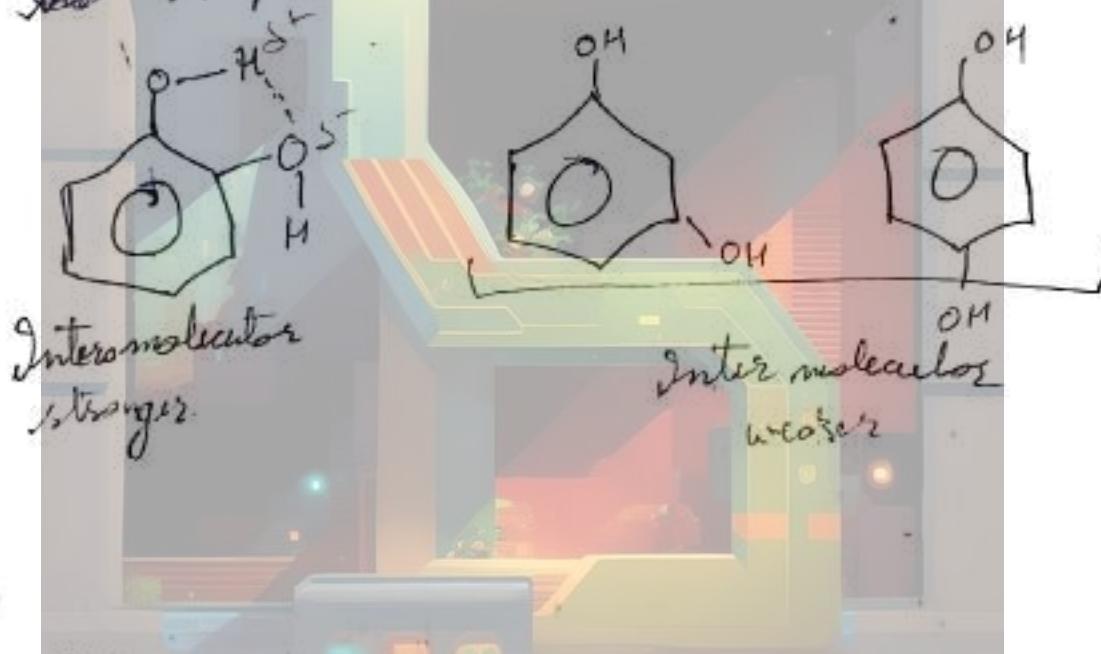
④ O^2 -hydroxy Acetone aldehyde
or
(Salicylaldehyde)





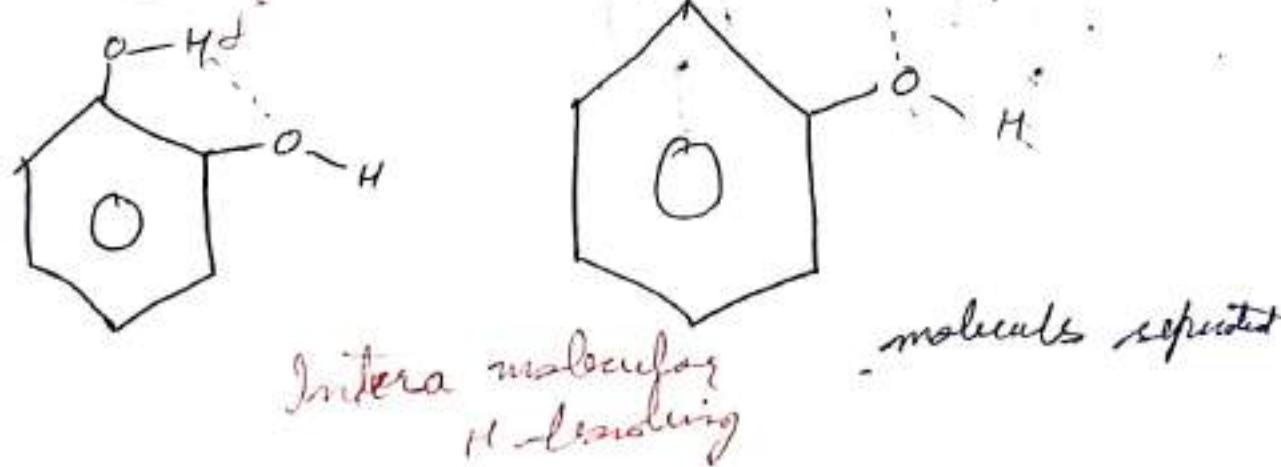
Note: - ① With respect to Benzene ortho position form intra molecular hydrogen bond with meta & para position can form inter molecular hydrogen bonding.

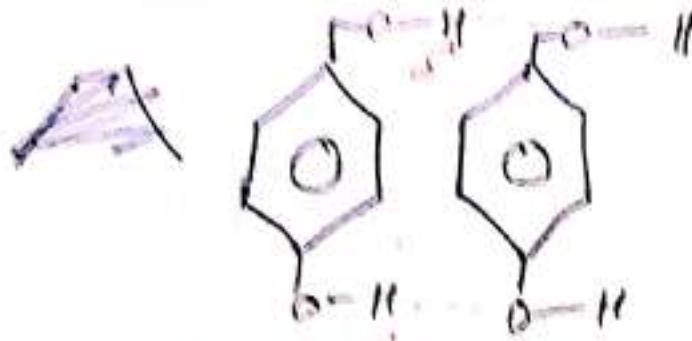
② Intra Molecular hydrogen bonding is more stronger than compared to inter molecular hydrogen bonding



Bonds

Boiling Point





Inter molecular
H-Bonding

molecules H-atoms are
connected by

Boiling Point \Rightarrow Intermolecular forces

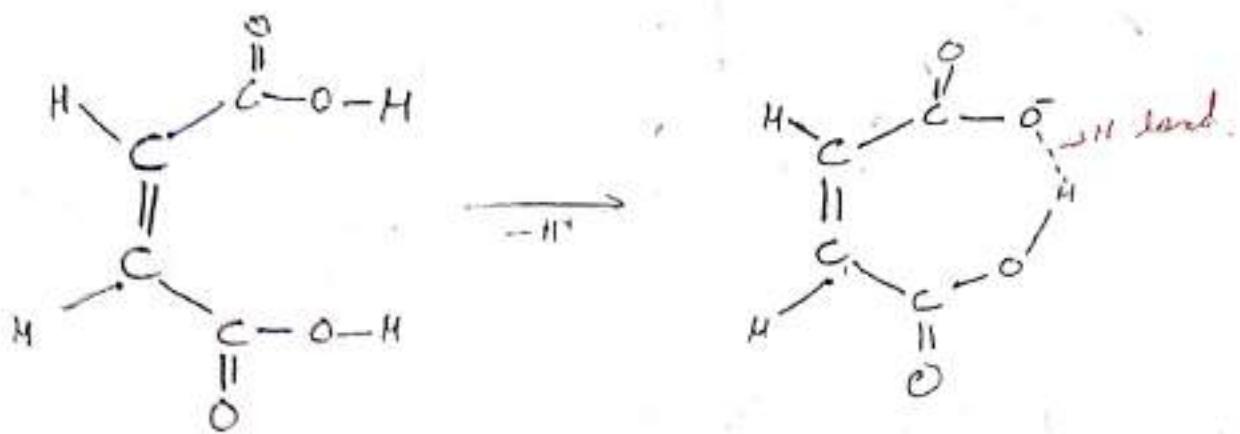
Volatility \Rightarrow vaporize more easily

Viscosity \Rightarrow restriction in movement of molecule
 $\text{V}_{\text{inter}} > \text{V}_{\text{intra}}$

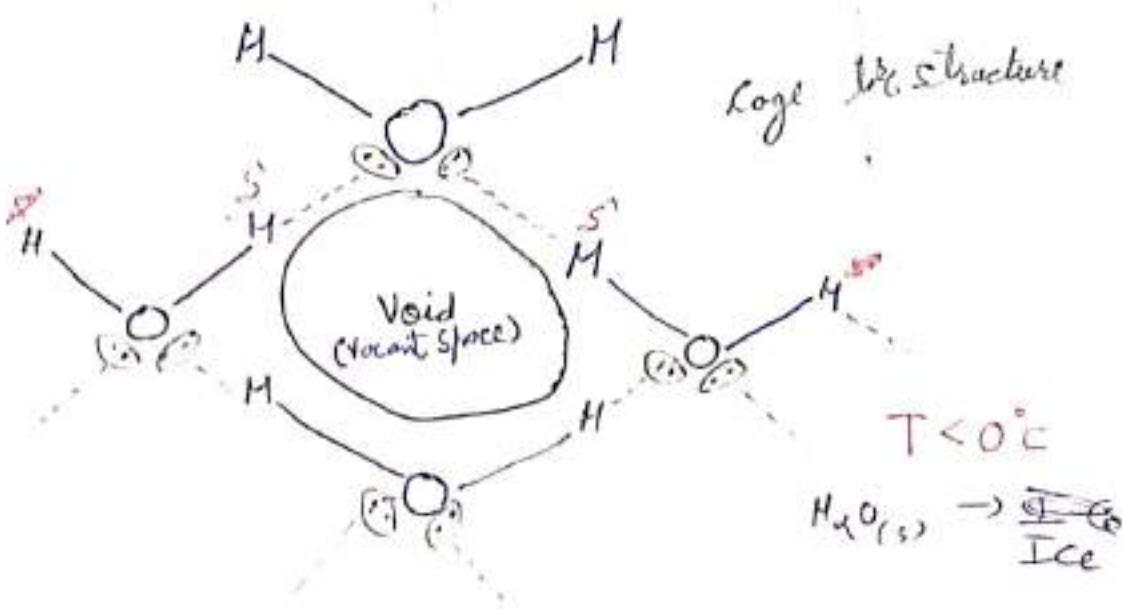
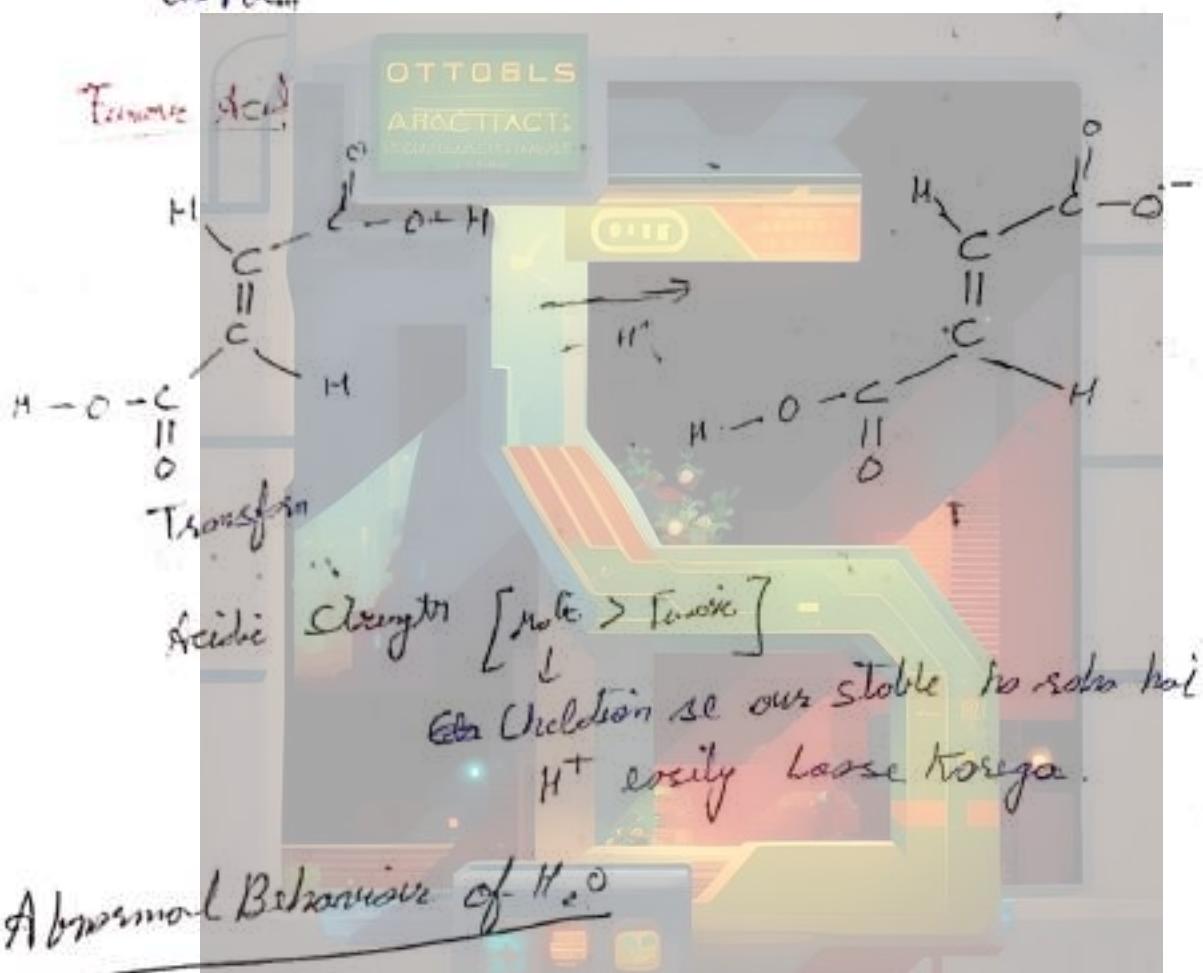
Solubility \Rightarrow less water i.e. solutes have less H-bond donor
 $\text{V}_{\text{inter}} > \text{V}_{\text{intra}}$



Q) Molar Acid



Cis Form



- In Ice Every oxygen atom is surrounded by 4 hydrogen atoms (cage like structure). In which 2 hydrogen atoms are bonded with covalent bonds & rest two hydrogen atoms are bonded with hydrogen bonding.
- Due to Intermolecular hydrogen bonding, voids are created in ice hence, volume increases.

$$d = \frac{\text{mass}}{\text{Volume}} \rightarrow \begin{matrix} \text{fixed} \\ \uparrow \end{matrix}$$

Volume ↑ ; Density ↓ / Ice. $\text{Temp} < 0^\circ\text{C}$

ii) For ~~Temp~~ $0^\circ\text{C} < \text{Temp} < 4^\circ\text{C}$

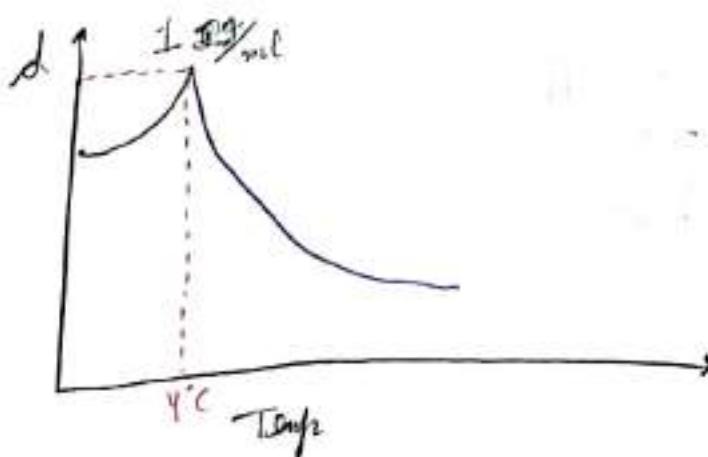
→ As Temperature Increases from 0 to 4°C Then voids created by hydrogen bonding are destroyed.

Volume ↓ ; Density ↑ Temp ($0 - 4^\circ\text{C}$)

iii) $\text{Temp} > 4^\circ\text{C}$ (Thermal Expansion)

→ As Temperature increases from 4°C & above, more no. of hydrogen bonds will break hence, volume increases.

Volume ↑ ; density ↓ Temp $> 4^\circ\text{C}$



Ques.:- ① Two ice cubes held together (force is applied) stick to each other due to hydrogen Bonding

②

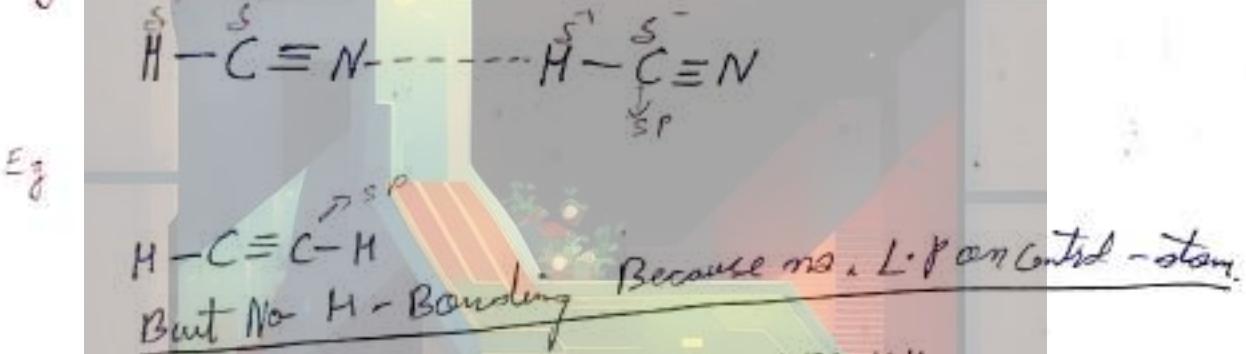
Q Why wood sticks are used to hold ice-cream.

~~Ans. Wood Stick~~

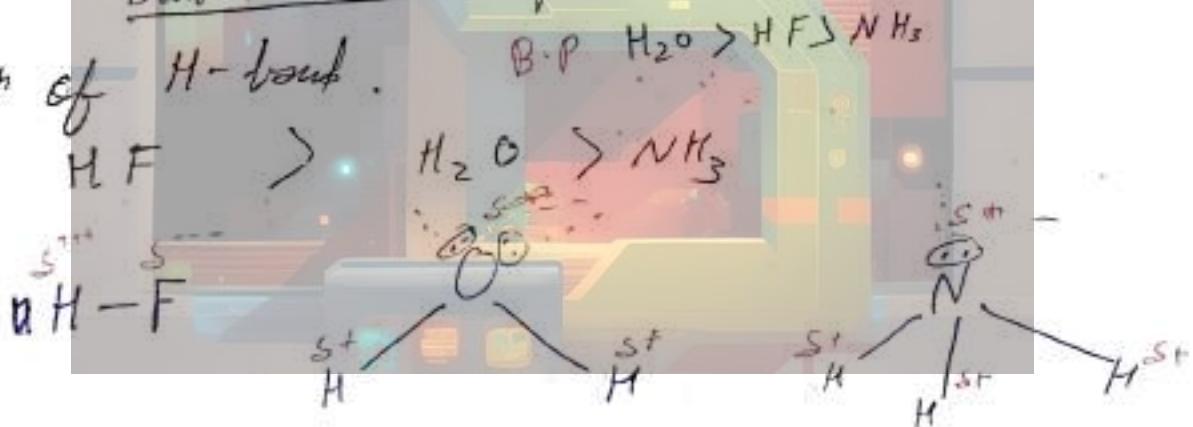
Ans. Wood stick contains cellulose which forms hydrogen bonding with water molecules in ice-cream.

Note:- ① Hydrogen attached with sp hybridised carbon may show hydrogen bonding.

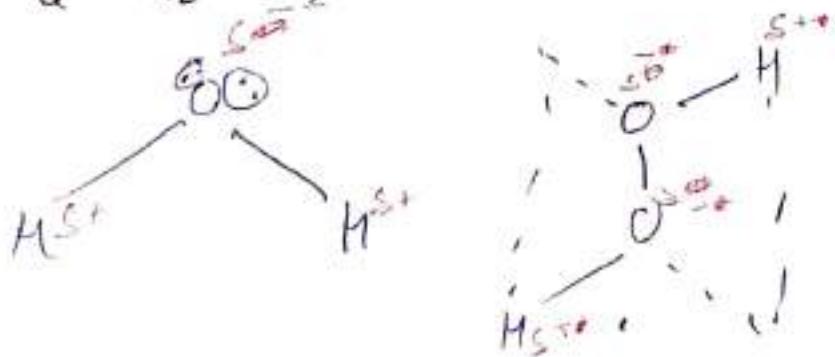
Eg



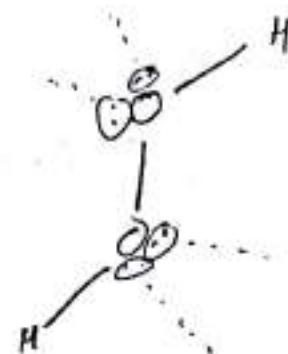
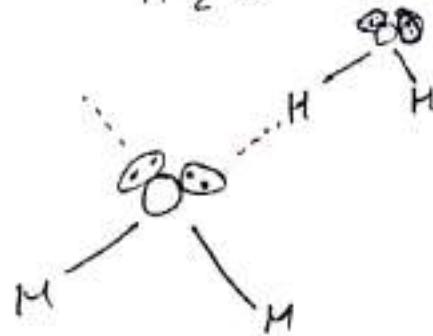
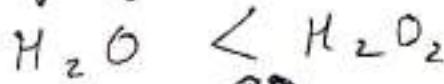
Q Strength of H-bond.



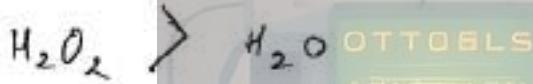
Q $\text{H}_2\text{O} > \text{N}_2\text{O}_2$



Q extent of hydrogen bonding (no. of H-Bond)



B.P (Boiling Point)



OTTOBL'S
ARCTACTIC
CONCENTRATOR

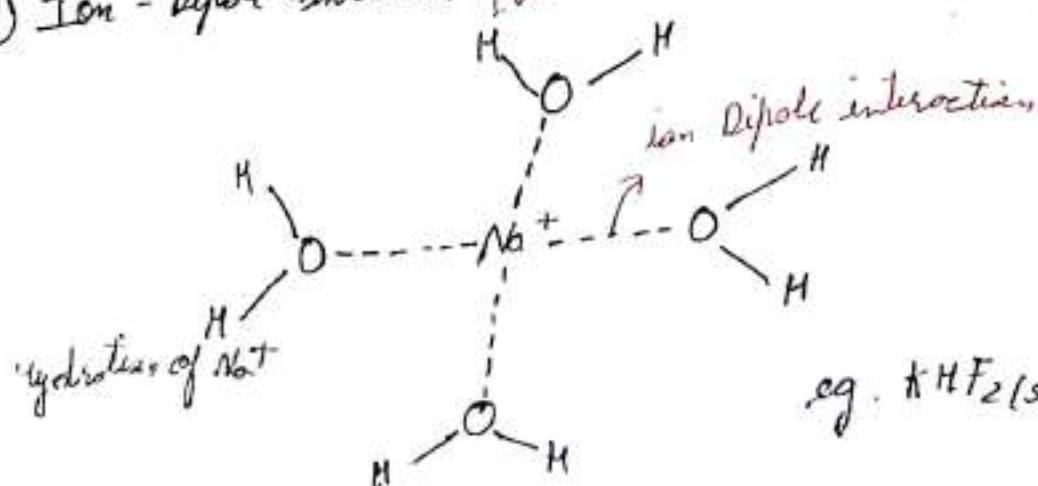
Weak Forces

involving ions

- ① → ion-dipole Interaction
- ② → ion-induced dipole interaction

- ① → no ions involved.
(Vander wall forces).
- ② → Dipole-Dipole Interaction
(Keesom force)
- ③ → Dipole-Induced ~~Dipole~~ Dipole
interaction (Debye-Force)
- ④ → Instantaneous dipole-induced
dipole (London Dispersion
Force) (L.D.F)

① Ion - Dipole Interaction (Interaction between ion & Polar molecule)

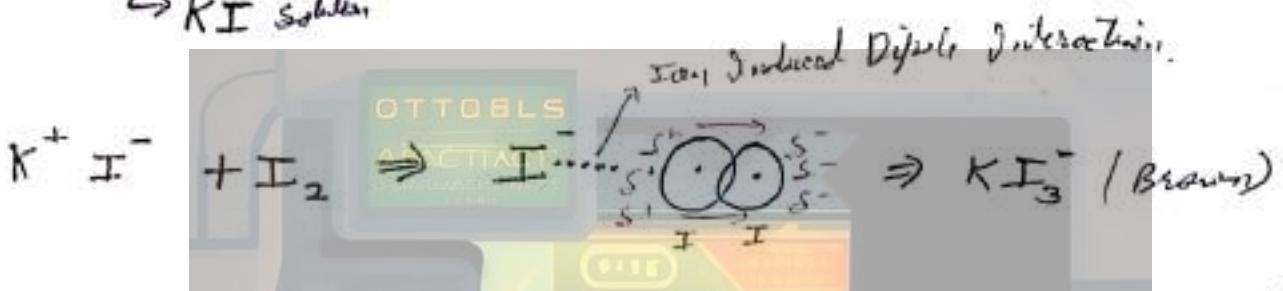
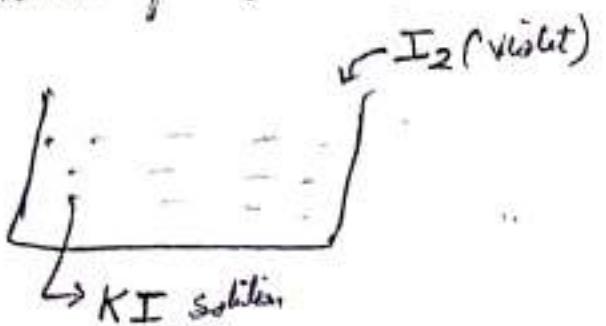


e.g. $\text{KHF}_2(s)$, $\text{KHCO}_3(s)$

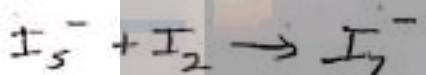
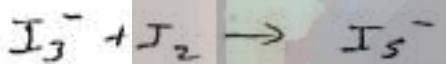
(2) Ion Induced Dipole Interaction

→ It is interaction between ion & non-polar molecule.

Eg. Formation of Poly Iodide ions.



→ ~~I-~~ I⁻ ke one charge ke reason saara O⁻ve charge dure side shift hoga mtl d ~~O⁻~~ve charge I⁻ ke O⁻ve se mix hoga bana leta hoga.



Q Stability of



Size ki major

③ Dipole-Dipole Interaction :- (Keeson-Force)

→ Interaction Between Polar Molecules



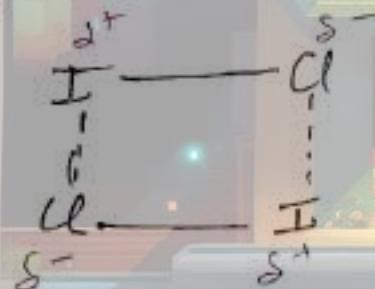
Note:- Hydrogen-Bonding is a special case of Dipole-Dipole Interaction

→ Two Types of arrangement

i) Head-Tail

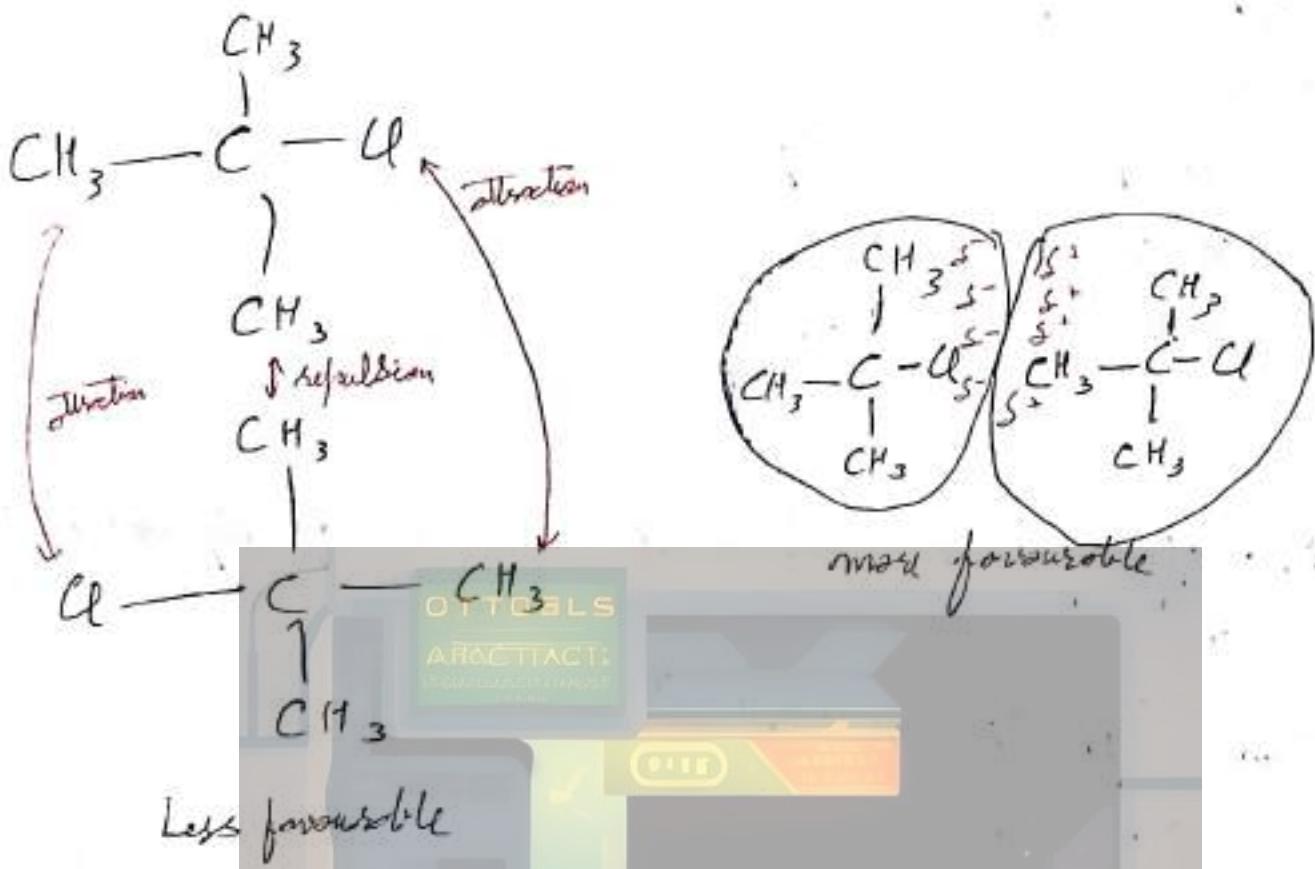


ii) Anti Parallel Arrangement (more stable)



→ Anti Parallel is more stable than Head-Tail.

→ In General Anti-Parallel Arrangement will be more stable as compared to Head-Tail Arrangement but for bulkier molecule anti-parallel arrangement is less favourable due to ~~steric~~ steric crowding.



H.W.

25-06-2024

O-1 (31-38, 41, 51, 52, 53)

O-2 (19-23)

S-1 (13)

S-2 (11, 13-18) (29)

J-M (3, 22, 28, 30, 38, 43, 44, 48, 51, 52, 54, 55)

J-A (5, 8, 13, 14, 15, 16, 11, 21)

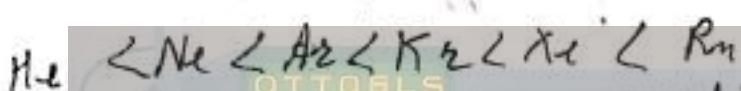
④ Dipole Induced Dipole Interaction (Debye Force)

→ Interaction between polar & non-polar molecule.

Eg. Solubility of inert gas.



Solubility order



Size ↑ \propto of H_2O ↑ Solubility ↑

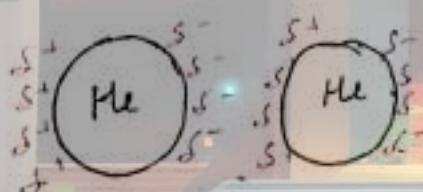
→ Polar nature of H_2O induce the charge shift in He.

⑤ Instantaneous Dipole induced Dipole Interaction

(London or Dispersion forces) (LDF)

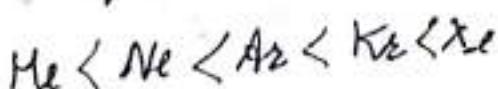
→ Interaction between two non-polar molecules.

Eg. formation of liquid helium



→ Fraction of second life e^- unsymmetrical have, instant dipole generate thus juste due He mol dipole induce bays.

Liquification of gases



Factors on which weak forces depend

weak forces \propto

- ① Surface area
- ② Molecular volume (size)
- ③ Molecular weight
- ④ No. of Polarizable e^-
- ⑤ $\frac{1}{\text{Temperature}}$

Energy Distance ~~Relation~~ for diff type of interaction.

① Ion - Ion interaction

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

② Ion - Dipole Interaction

$$E \propto \frac{1}{r^2}$$

③ Dipole - Dipole Interaction

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

④ Ion-induced dipole

$$E \propto \frac{1}{r^4}$$

⑤ Dipole induced dipole

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

⑥ Instantaneous dipole induced dipole

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

(r is Intermolecular Distance)

Strength of weak forces.

(Ion - Ion) $>$ (Ion-Dipole) $>$ (Dipole-Dipole) \checkmark

(Instantaneous dipole induced dipole) $<$ (dipole induced dipole) $<$ (ion induced dipole)

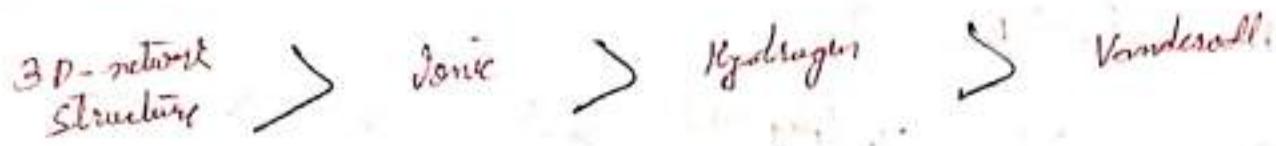
Note:- ① London Dispersion Force exist everywhere except ideal gas.

- ② between the polar molecule 3 forces exist
1. Dipole - Dipole
 2. Dipole induced Dipole
 3. LDF

③ No bond is $\% 100$ covalent due to LDF.

Boiling Point Order

In General



Boiling Point order.

① $H_2 < D_2 < T_2$

molecular weight \uparrow Work force \uparrow $\therefore B.P \uparrow$

② $BF_3 < BCl_3 < BBz_3 < BI_3$

③ $BF_3 < B(OCH_3)_3 < B(OCH_2CH_3)_3$

④ $B(OH)_3 > BF_3$ (H-bond).

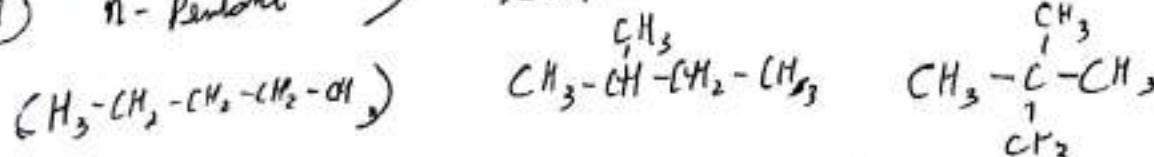
⑤ $B(OH)_3 > B(OCH_3)_3$ (H -bond)

⑥ $CH_4 < SiH_4 < GeH_4 < SnH_4$ (m. weight)

⑦ $R-O-R < R-SR$ (R -Alkyl group) (m. weight)

⑧ $I-Cl > Br_2$ (L.P.F.)

⑨ $n\text{-Pentane} > iso\text{-pentane} > Neo\text{-pentane}$ (surface area)



Note:- Branching ~~decreases~~ increases Surface area

⑩ $\text{CH}_4 < \text{CF}_4 < \text{CCl}_4$ (M. weight)

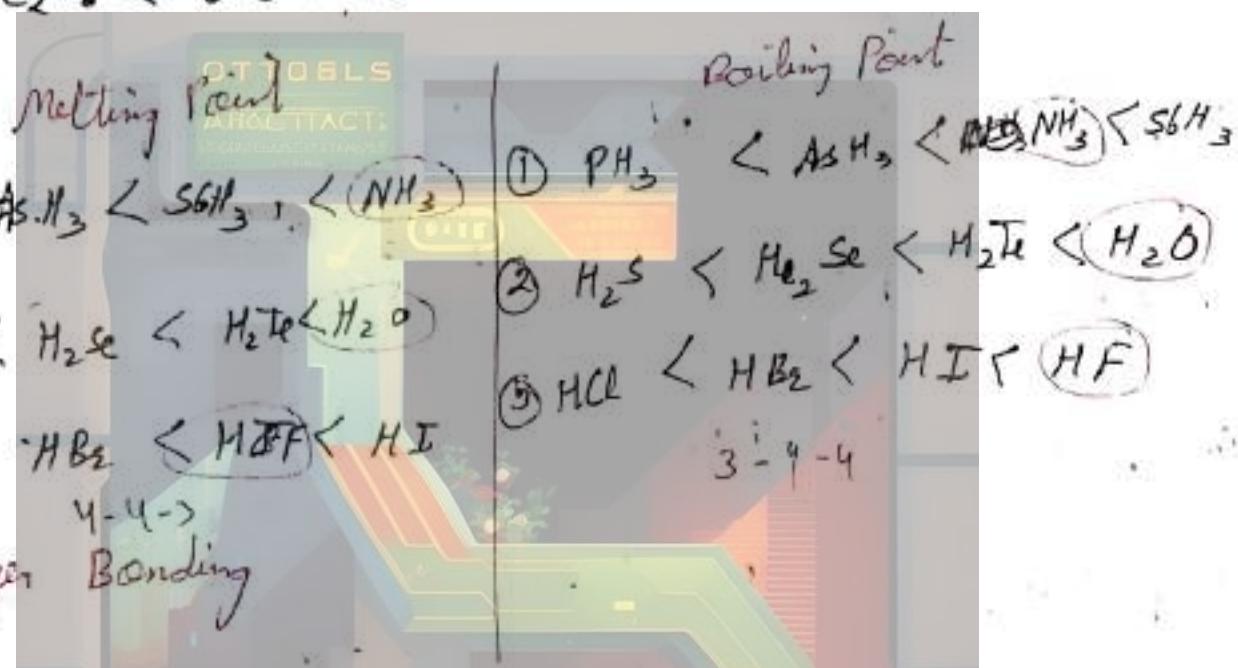
⑪ $\text{NF}_3 < \text{N}(\text{CH}_3)_3$ (A. weight)

⑫ $\text{F}_2 < \text{Cl}_2 < \text{Br}_2 < \text{I}_2$ (M. weight)

⑬ NaCl SiO_2 CO_2 (dry ice) $\text{SiO}_2 > \text{NaCl} > \text{CO}_2$
(ionic) (P-P) (P-P)

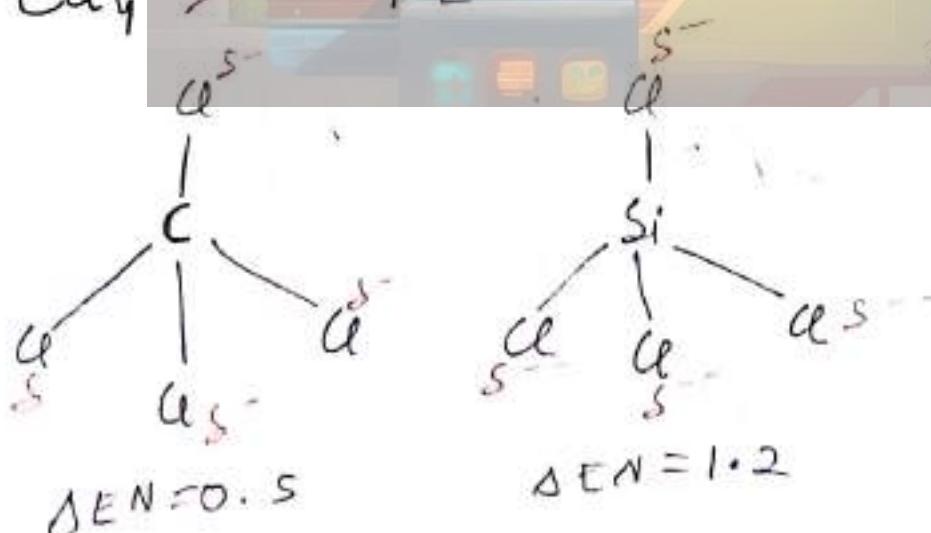
⑭ $\text{He} < \text{H}_2 < \text{D}_2 < \text{I}_2$

⑮ $\text{CH}_4 < \text{C}_2\text{H}_6 < \text{C}_3\text{H}_8 \dots$ (S. over, M. weight)



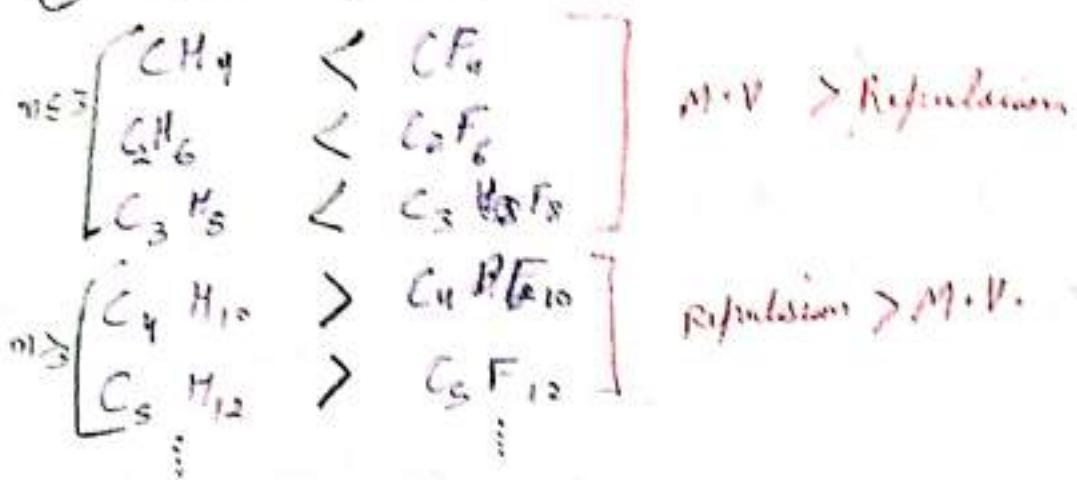
⑯ $\text{Ca} < \text{SiCl}_4$ (Expected)

$\text{CCl}_4 > \text{SiCl}_4$ [Actual]



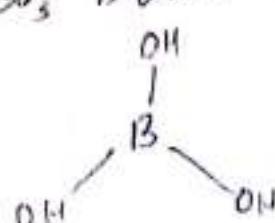
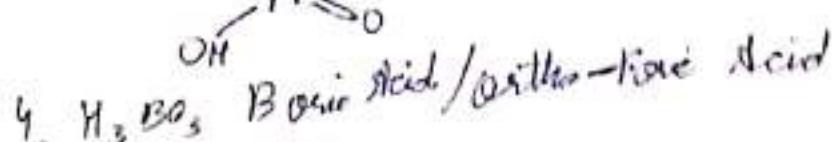
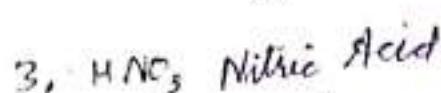
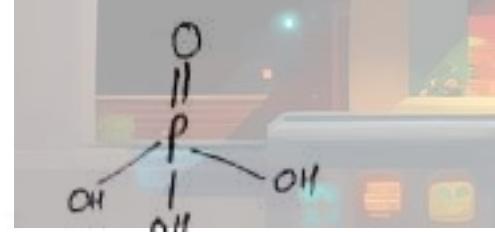
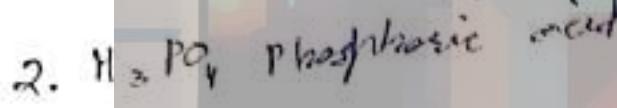
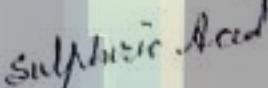
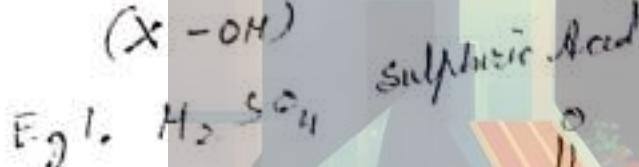
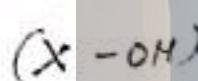
more charge on periphery of
 SiCl_4 , repulsion ↑;
 $\text{B.P.} \downarrow$

17 Alkane (C_nH_{2n+2})



Oxy Acids

- Acid which contains Oxygen is known as oxy acid.
- Acid which OH- group is directly bonded with non-metal except Hydrogen.



→ Always use 'ic' suffix in G-13 & G-14 acids

G-13 - H_3BO_3 Boric Acid

G-14 - H_2CO_3 carbonic Acid

H_4SiO_4 Ortho Silicic Acid

G-15 - $HNO_3 \rightarrow$ Nitric Acid

$HNO_2 \rightarrow$ Nitrous Acid

$PH_3PO_4 \rightarrow$ Phosphoric Acid / ortho Phosphoric acid

H_3PO_3 $\xrightarrow{T \rightarrow 0}$ Phosphorous acid / ortho Phosphorous acid
→ oxidation state of central atom $\leq 3 \rightarrow$ 'aus'
 $> 3 \rightarrow$ 'ic'

G-16 $H_2SO_4 \rightarrow$ Sulphuric Acid

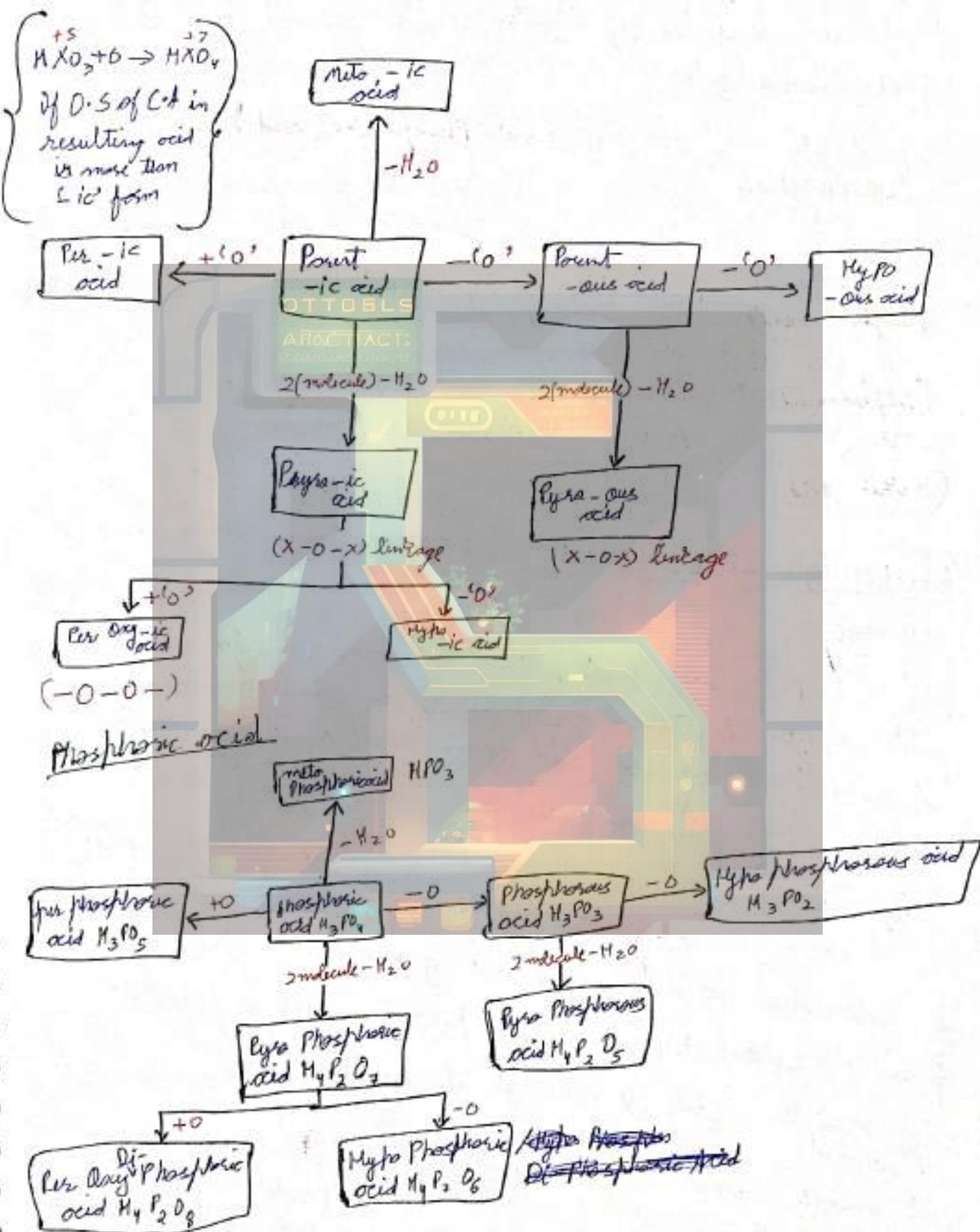
$H_2SO_3 \rightarrow$ Sulphurous Acid

→ oxidation state $\leq 4 \rightarrow$ 'aus'
 $> 4 \rightarrow$ 'ic'

G-17 $HClO_3 \rightarrow$ chloric acid
 $HClO_2 \rightarrow$ chlorous acid

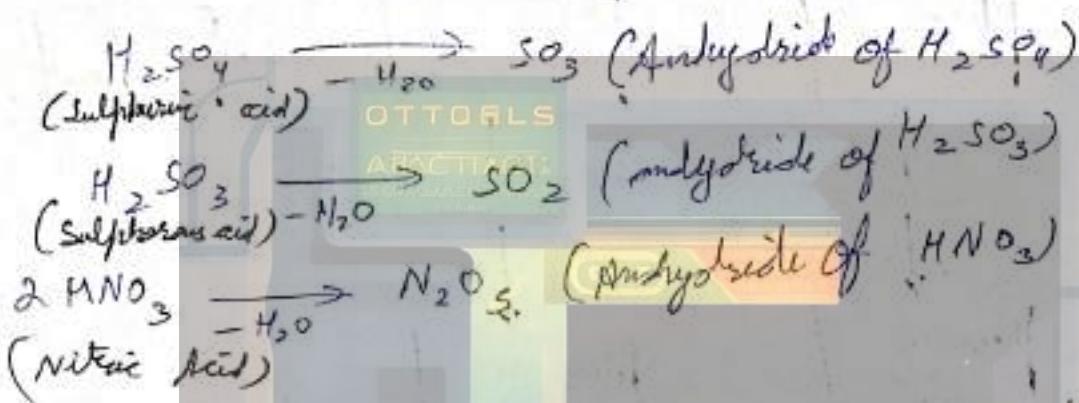
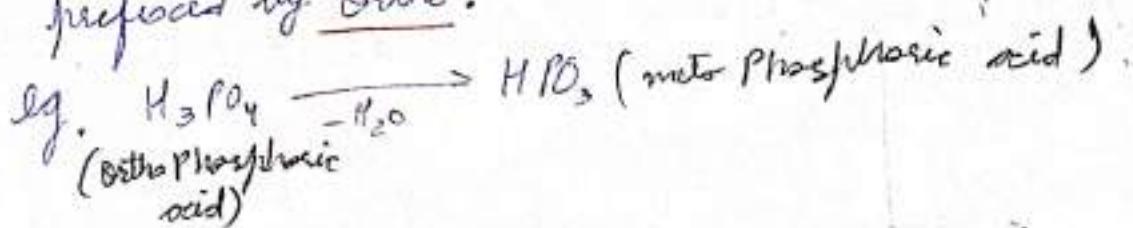
→ O-S- $\leq 3 \rightarrow$ 'aus'
 $> 3 \rightarrow$ 'ic'

Prefixes for oxy acids



General Structure of oxy-acids :-

- ① When we remove 1 H₂O molecule from one molecule of oxy-acid (X-OH) and resulting product is also oxy-acid (OX-OH) then product is prefixed by 'meta' and parent acid is prefixed by 'ortho'.



- ② Except Pyro Sulfuric acid other Pyros form having (X-O-X) linkage.

③ For S & P- Block element, their maximum possible oxidation state is equal to their group number in roman numerical and if element belongs to odd group number then oxidation states are common for it.

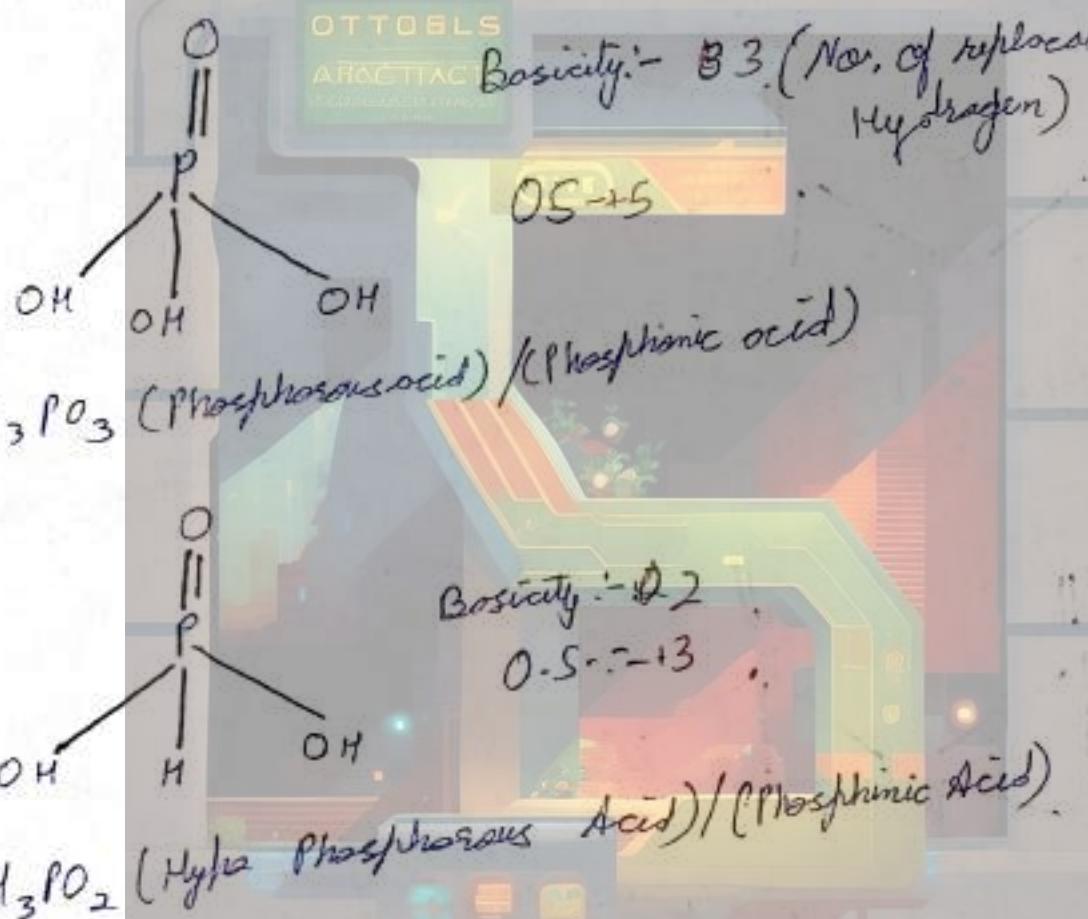
In case of even group number, even oxidation states are common.

- ④ If calculated oxidation state is greater than its maximum possible oxidation state it means peroxy linkage is present.

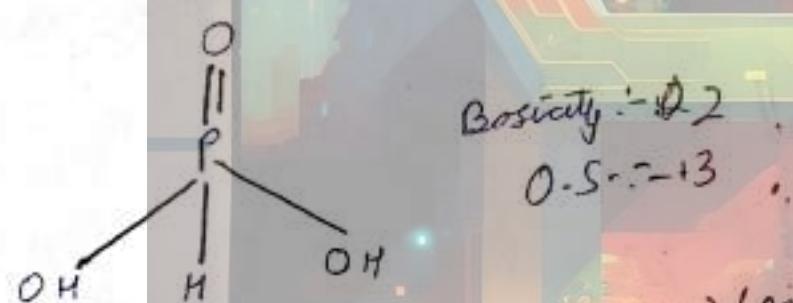
- ⑤ If calculated oxidation state is one of its common oxidation state it means oxy-linkage is present. (not valid for Pyro Sulfuric acid & Pyro This Sulfuric acid)

- ⑥ If oxidation state is un-common it means $(X-X)$ type linkage is present.
- ⑦ While drawing structure of oxy-acid first convert all hydrogen in OH group and attach symmetrically to central atom but in oxy-acid of phosphorous, first form $(P=O)$ then remaining oxygen are converted into OH group.
- Oxy acids of Phosphorus

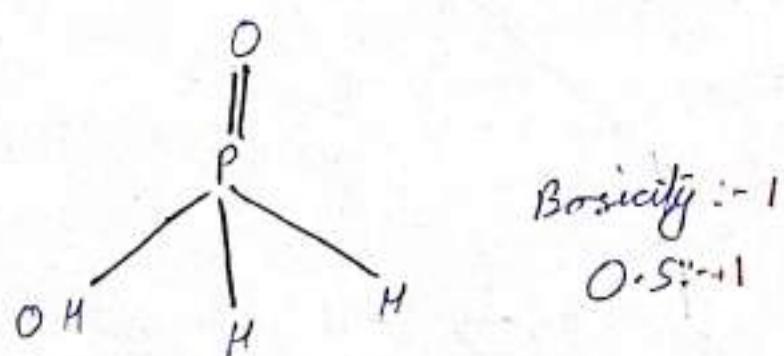
① H_3PO_4 (Ortho Phosphoric acid)



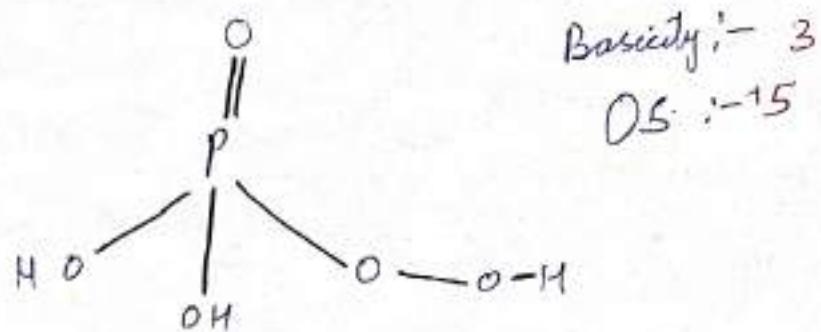
② H_3PO_3 (Phosphorous acid) / (Phosphinic acid)



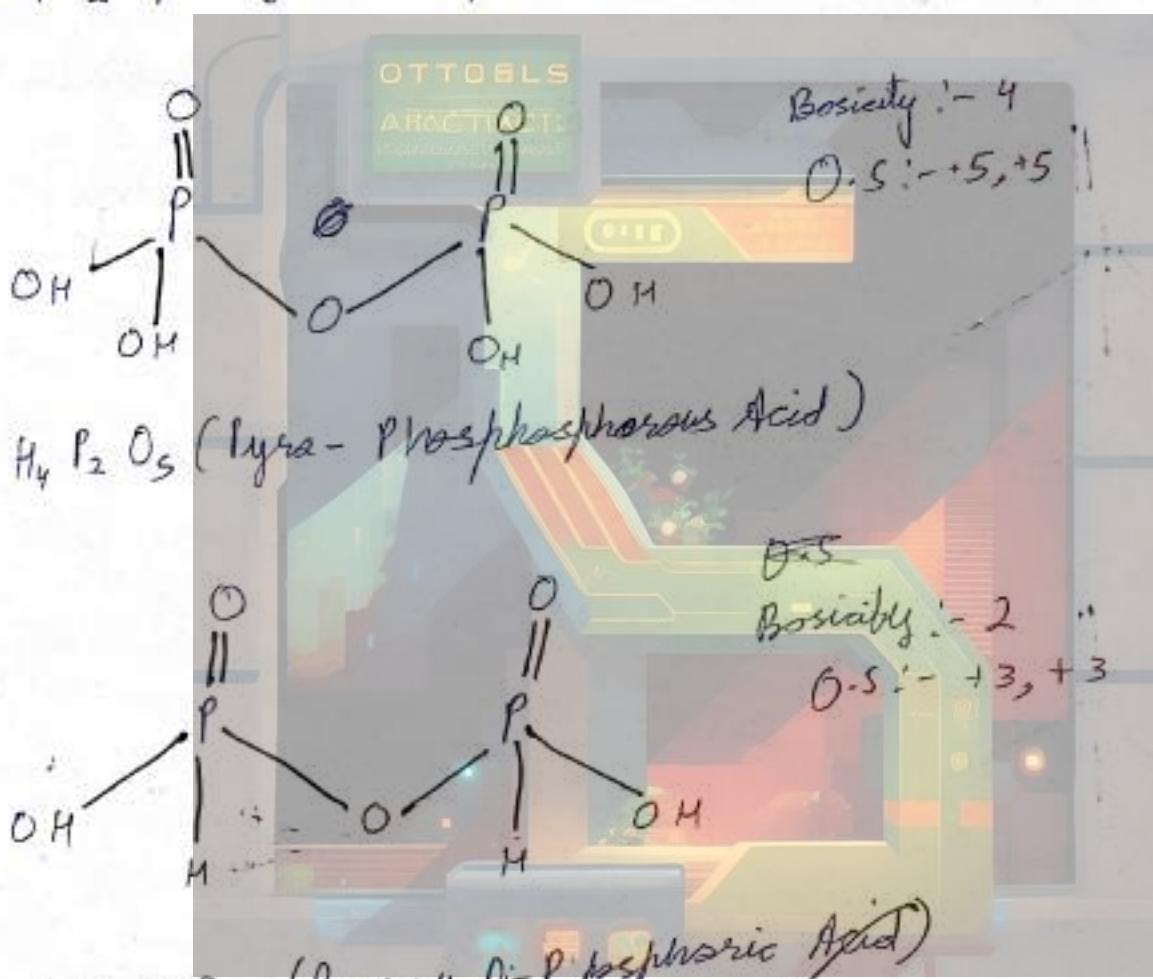
③ H_3PO_2 (Hydro Phosphorous Acid) / (Phosphinic Acid)



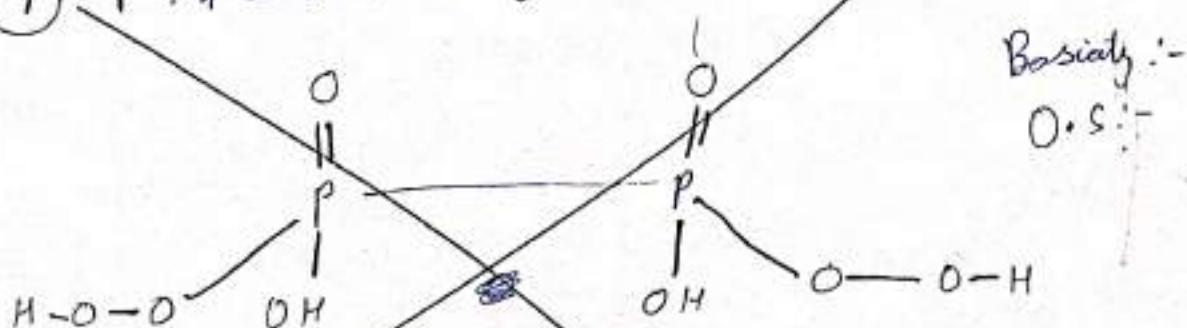
(4) H_3PO_5 (Per-Phosphoric acid) / (Peroxy mono phosphonic acid)



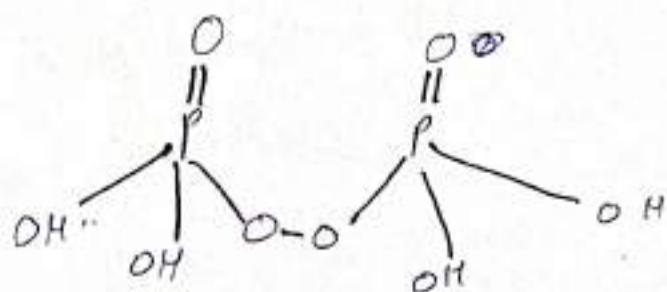
(5) $H_4P_2O_7$ (Pyro- Phosphoric acid)



(6) $H_4P_2O_5$ (Pyro- Phosphorous Acid)

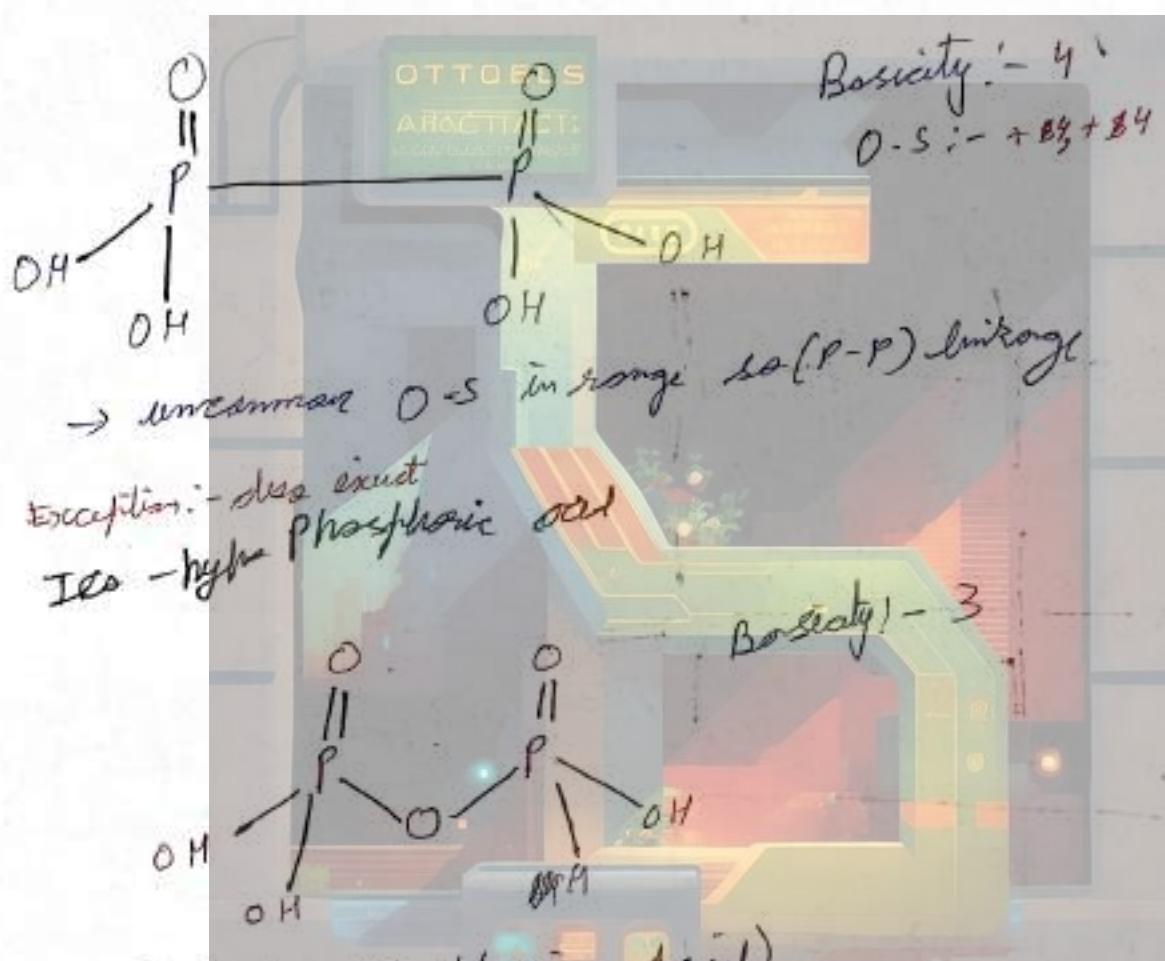


⑦ $H_2P_2O_8$ (Per oxy Di Phosphoric Acid)



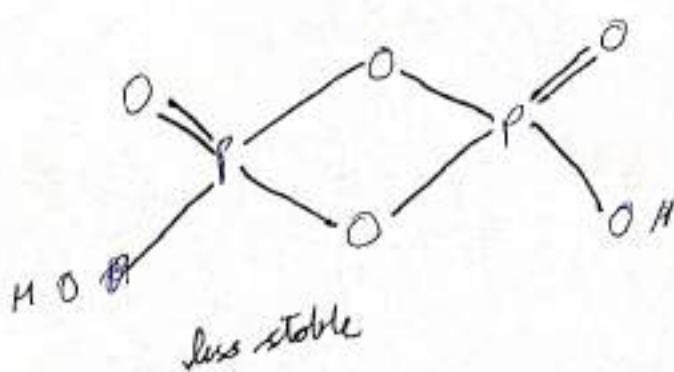
Basicity:- 4
O.S.: - +5, +5

⑧ $H_4P_2O_6$ (Hydro-Phosphoric Acid)

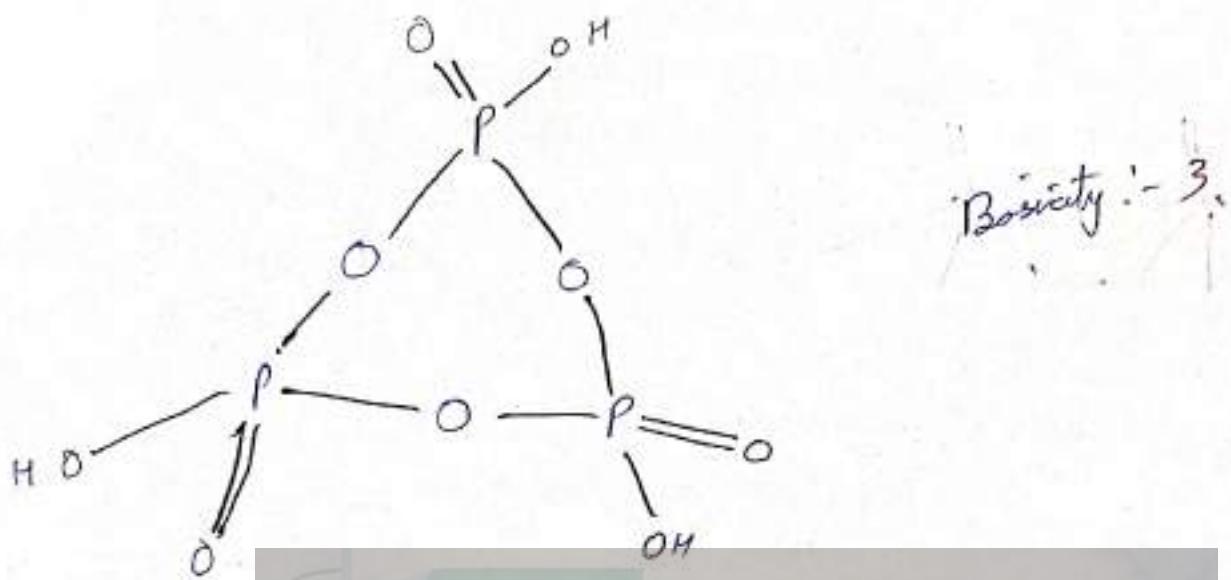


⑨ $(HPO_3)_n$ (Meta Phosphoric Acid)
Polymeric form

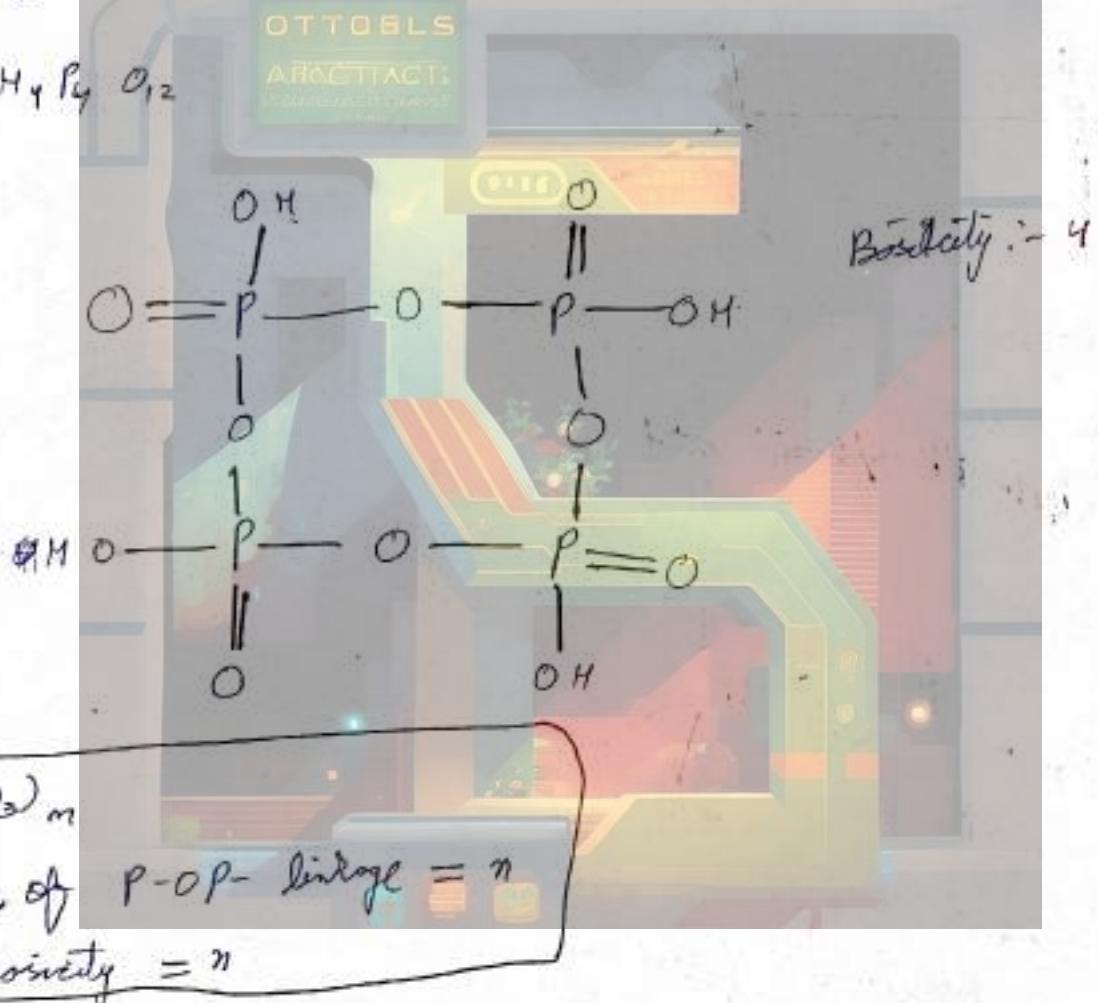
$n \geq 2, H_2P_2O_8$



$n=3, H_3P_3O_9$

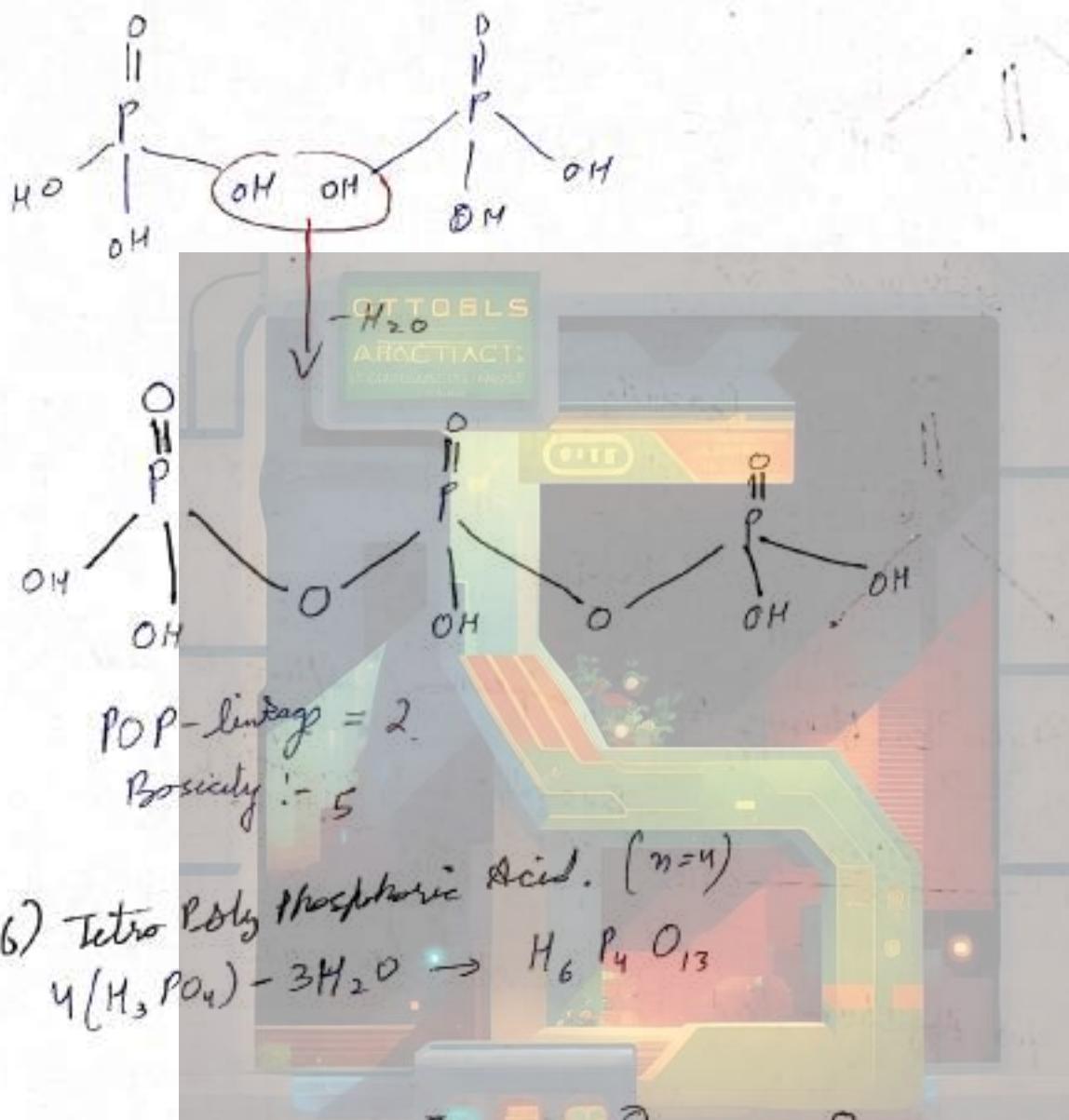
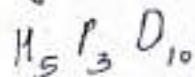


$n=4, H_4P_4O_{12}$

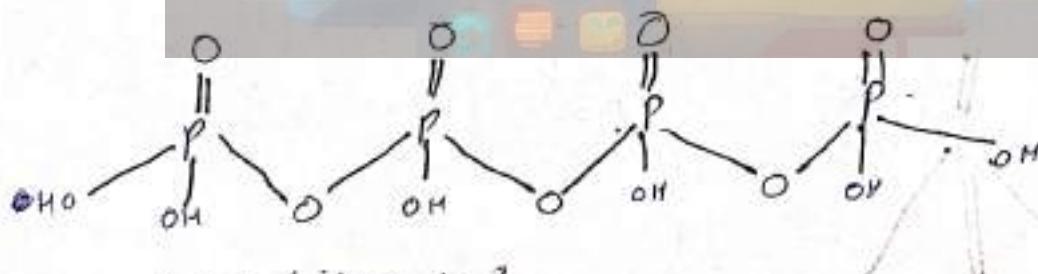
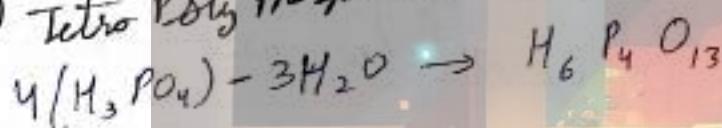


(10) Poly Phosphoric Acid: More than 2 phosphoric acid unit stuck to form chain like structure.

a) Tri Poly Phosphoric Acid. ($n=3$)



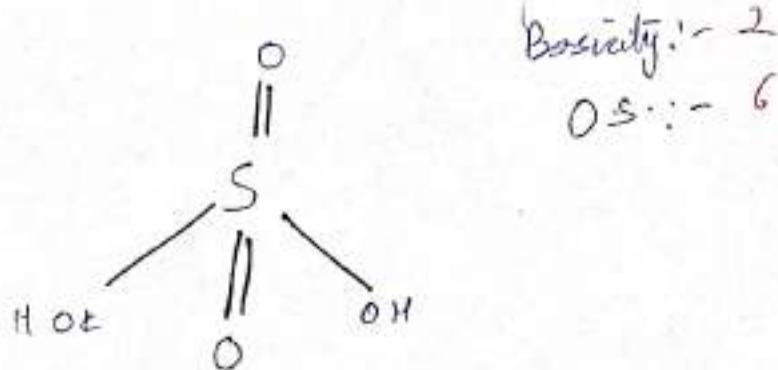
b) Tetra Poly Phosphoric Acid. ($n=4$)



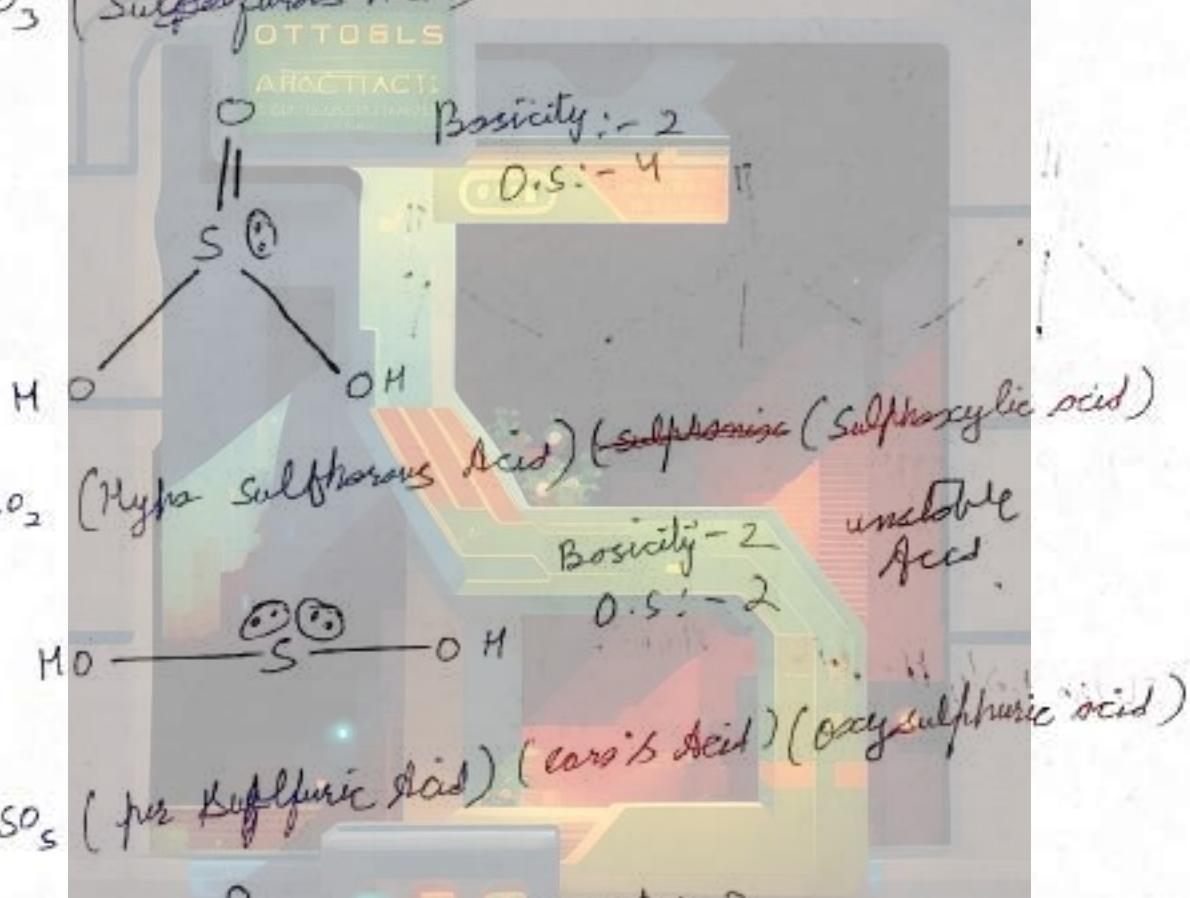
General: Basicity :- $n+2$
POP linkage :- $n-1$

Oxy Acids of Sulfur:-

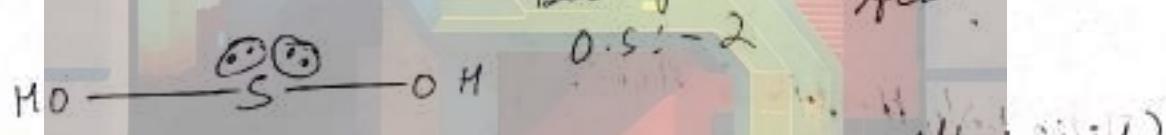
① H_2SO_4 (Sulfuric Acid)



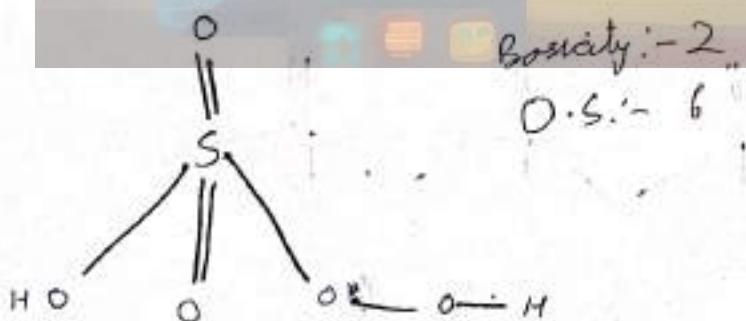
② H_2SO_3 (Sulfurous Acid)



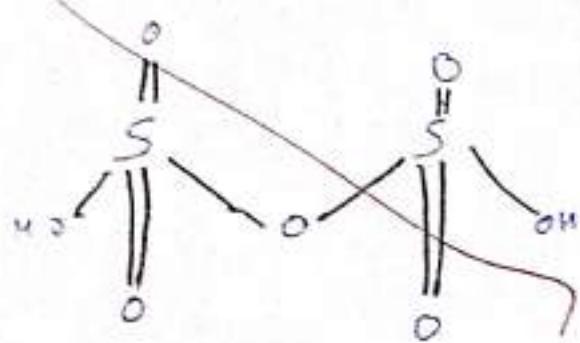
③ H_2SO_2 (Hydro sulfurous Acid)



④ H_2SO_5 (per Sulphuric Acid)



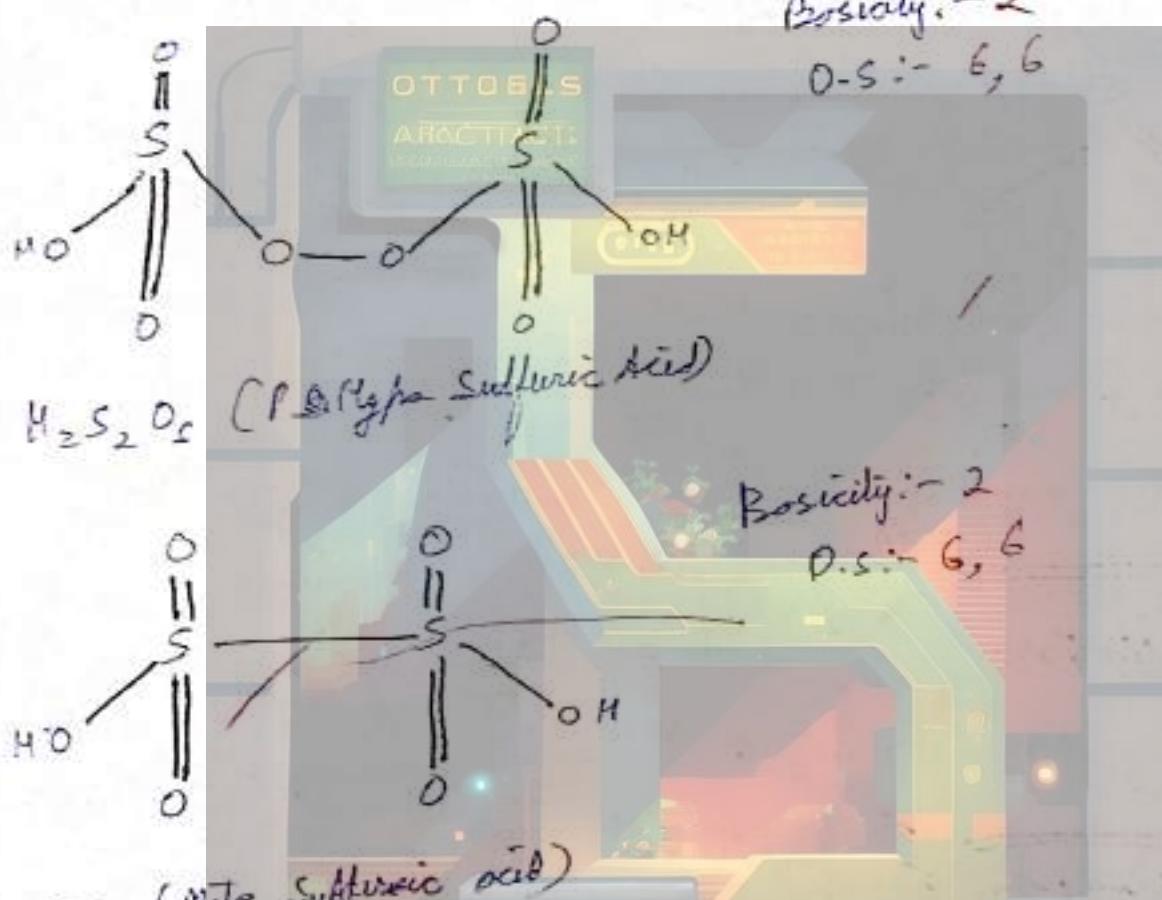
⑤ $H_2S_2O_7$ (Pyro Sulfuric Acid) (Oleum)



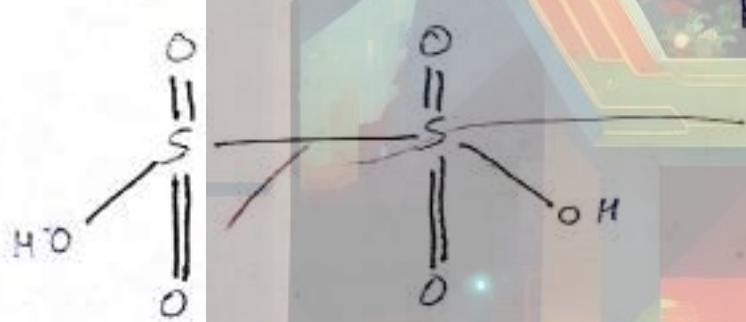
Basicity :- 2

O.S. :- 6, 6

⑥ $H_2S_2O_5$ (Iodoxy Di-Sulfuric Acid) (Mordell's Acid)



⑦ $H_2S_2O_5$ (Pyro Sulfuric Acid)



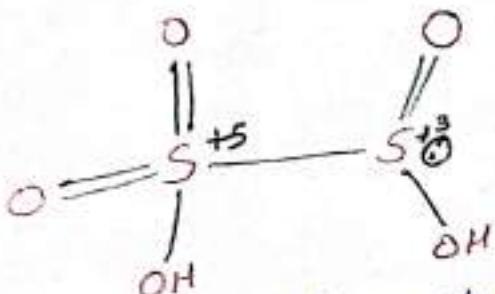
Basicity :- 2
O.S. :- 6, 6

⑧ SO_3 (Meta Sulfuric acid)
Not Exist

⑨ $H_2S_2O_5$ (Pyro Sulfurous Acid)

Basicity :- 2

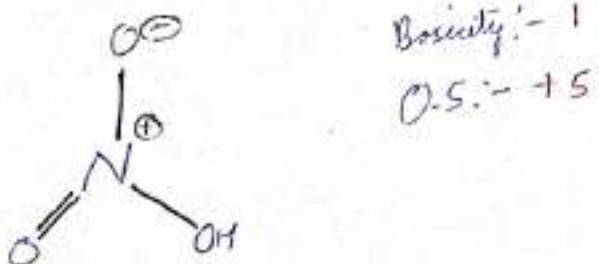
O.S. :- 5, 3



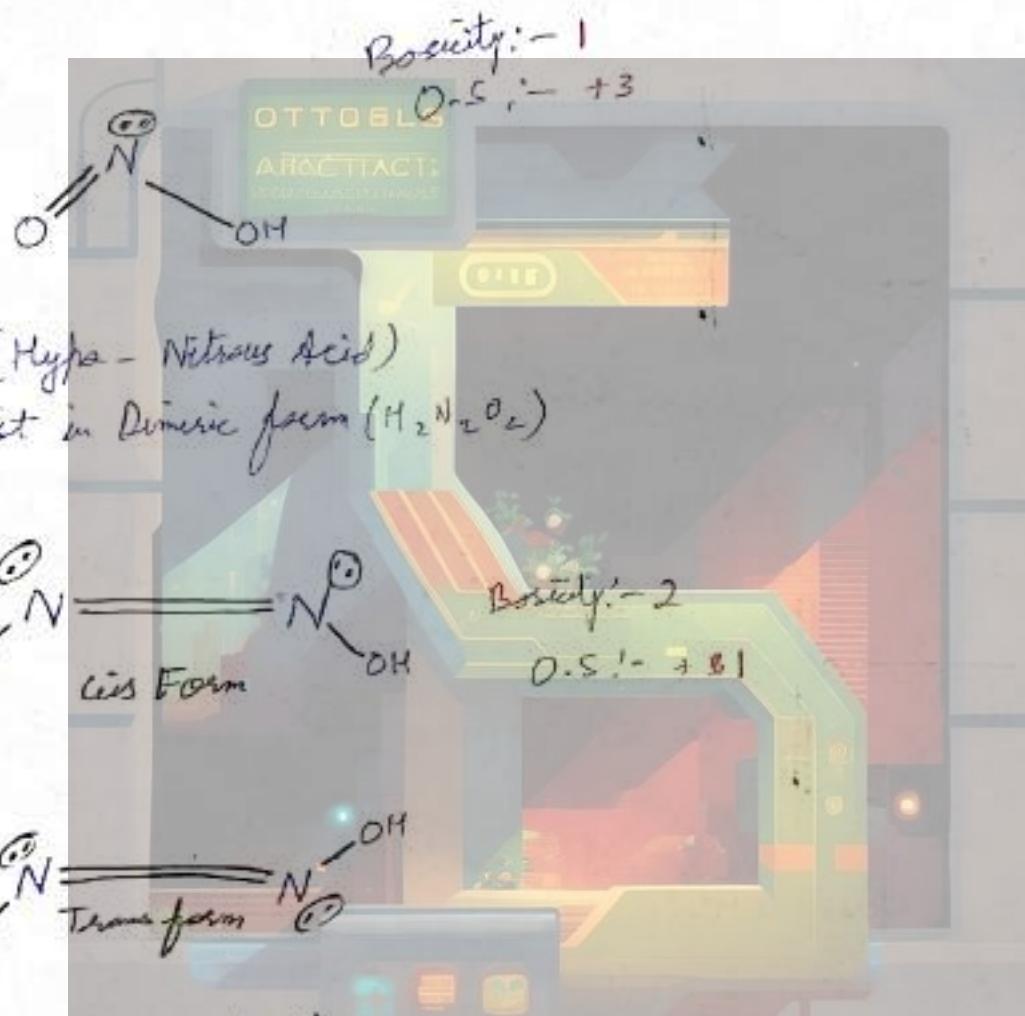
⑩ SO_2 (Meta Sulfurous Acid)
Not Exist.

Oxy Acids of Nitrogen

(1) HNO_3 (Nitric Acid)



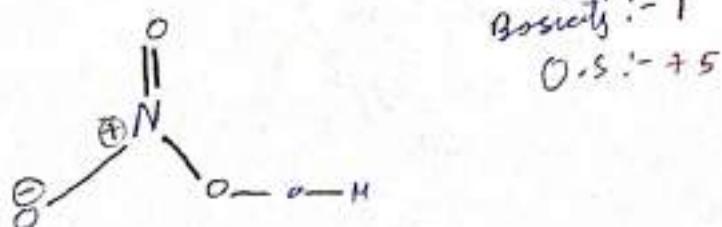
(2) HNO_2 (Nitrous Acid)

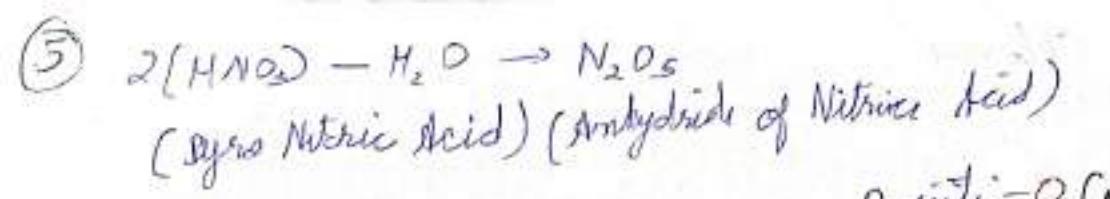


(3) HNO (Hydro-Nitrous Acid)

→ Exist in Dimeric form ($\text{H}_2\text{N}_2\text{O}_2$)

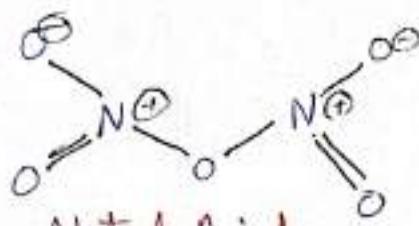
(4) HNO_4 (Per-Nitric Acid)



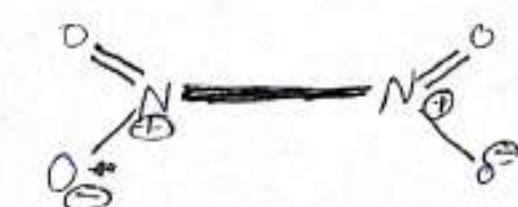
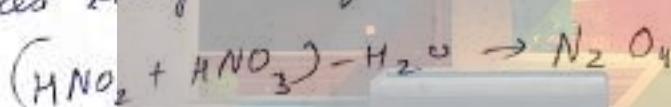
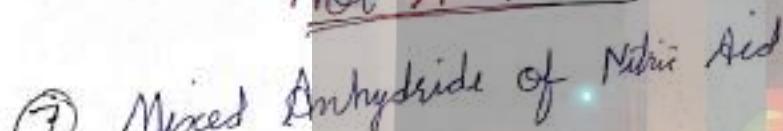
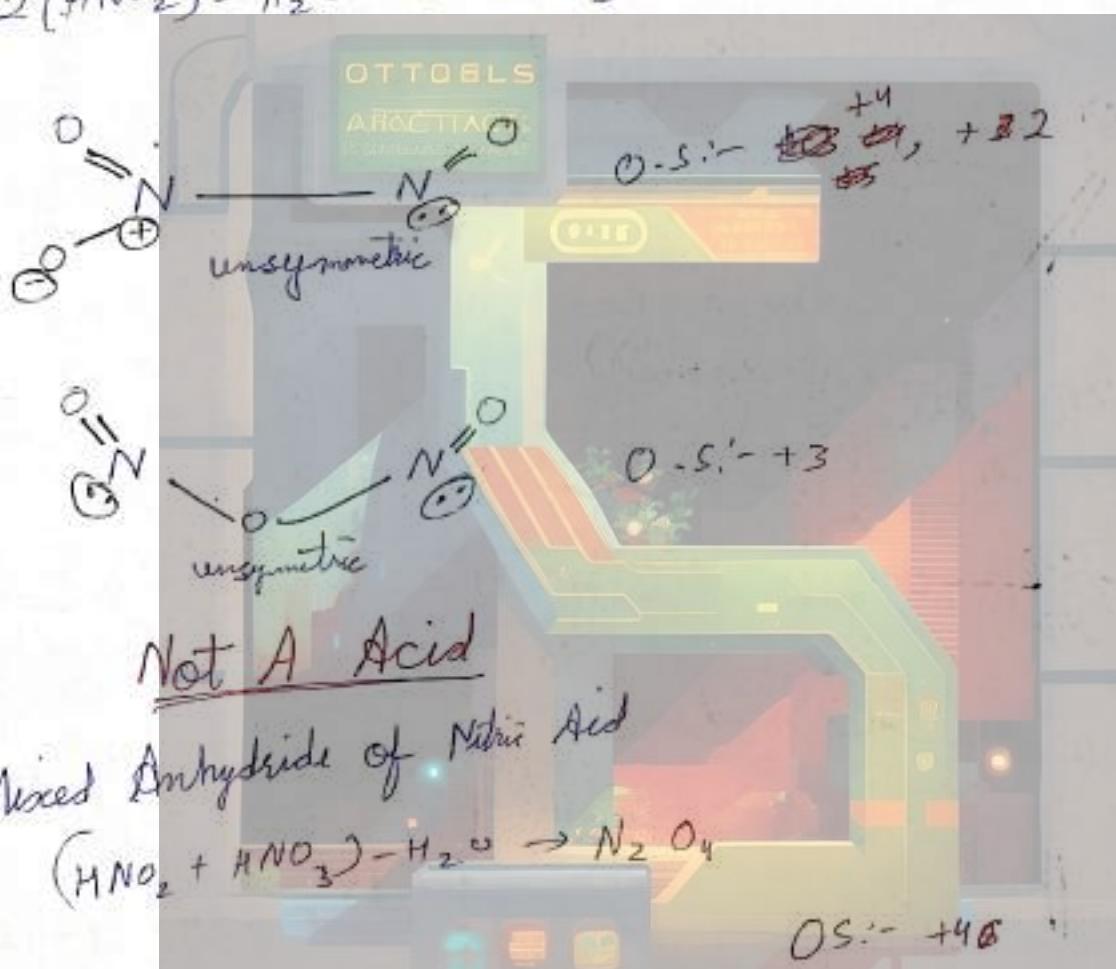
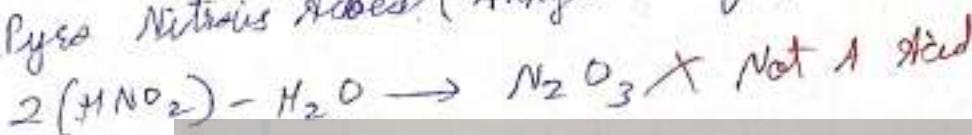
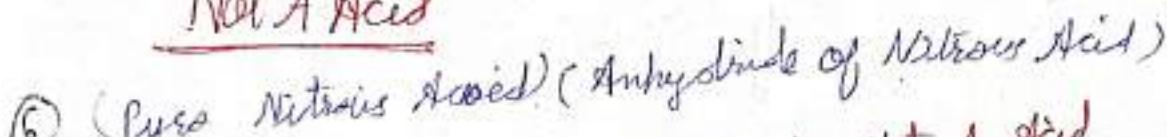


Basis of: -O (Not A Acid)

O.S.: -5



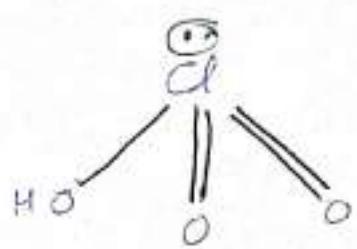
Not A Acid



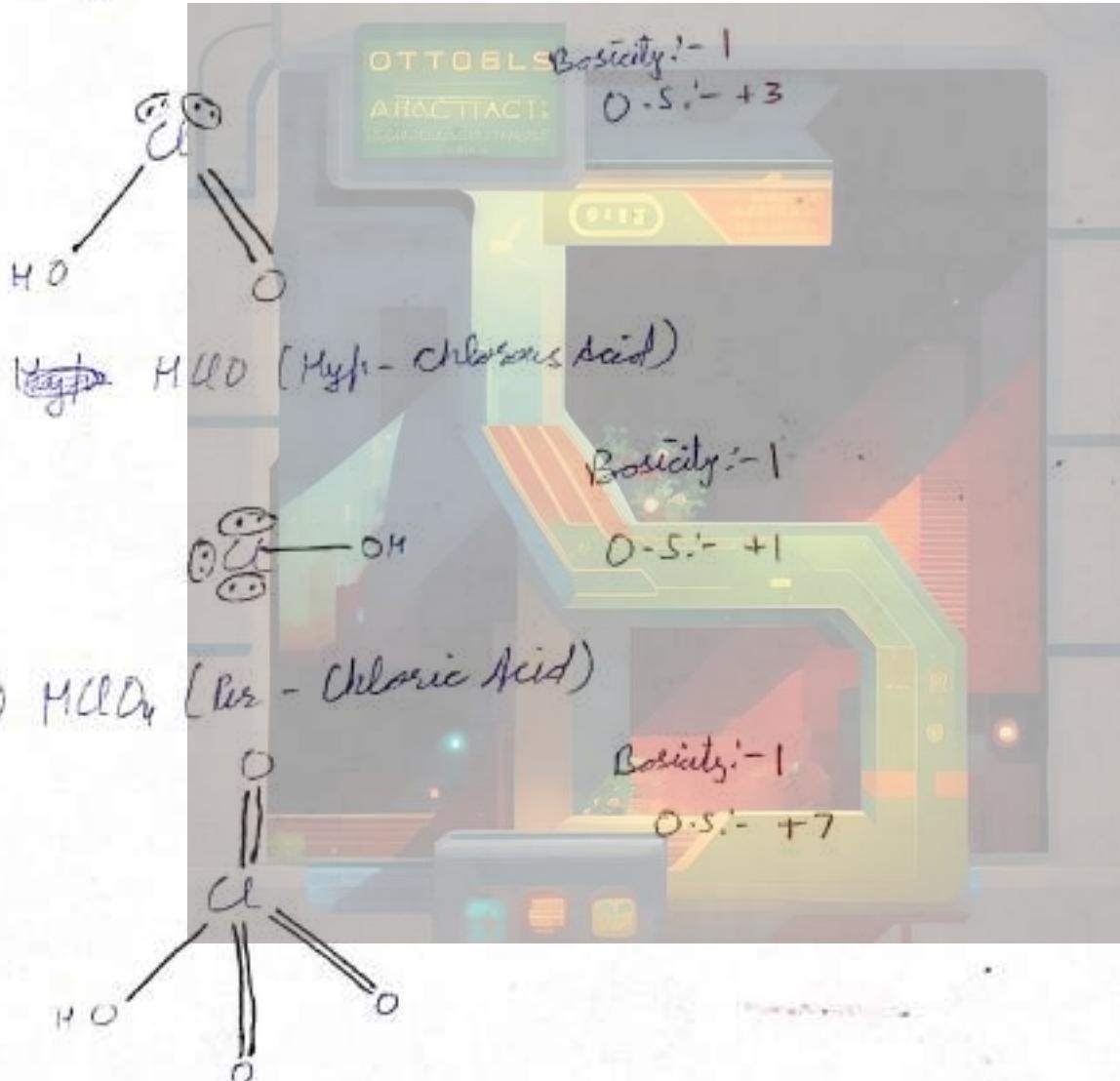
Not A Acid

Dioxy Acids of Chlorine

① HClO_3 (Chloric Acid)



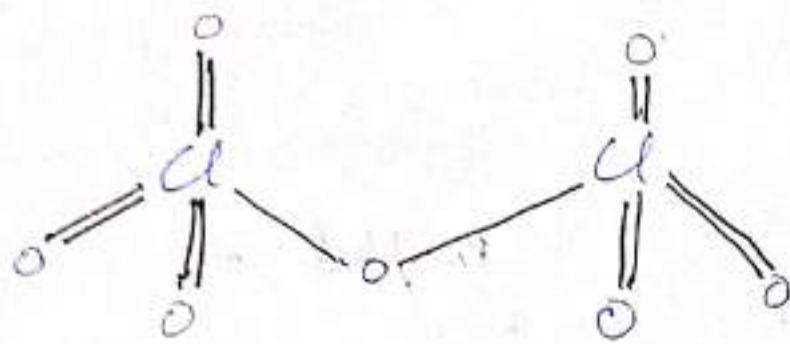
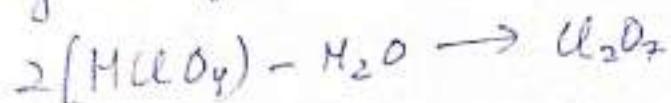
② HClO_2 (Chlorous Acid)



③ ~~HClO~~ HClO (Hypochlorous acid)

④ HClO_4 (Perchloric Acid)

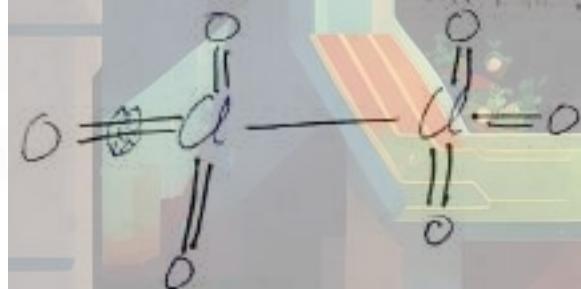
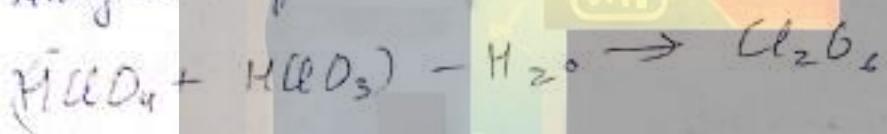
⑤ Anhydride of perchloric acid:



Basicity: - 0
O.S.: - 7

Not A Acid

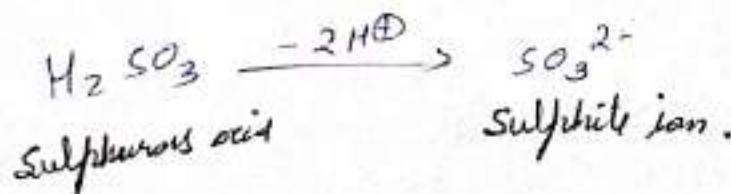
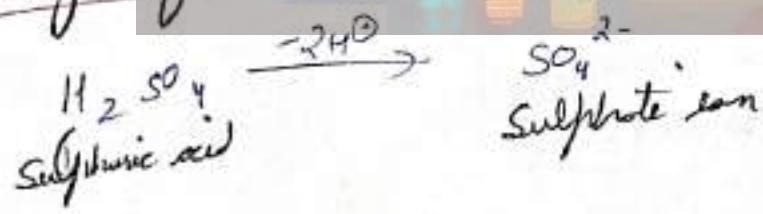
⑥ Anhydride of perchloric & chloric acid:-

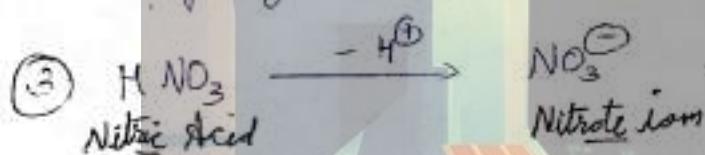
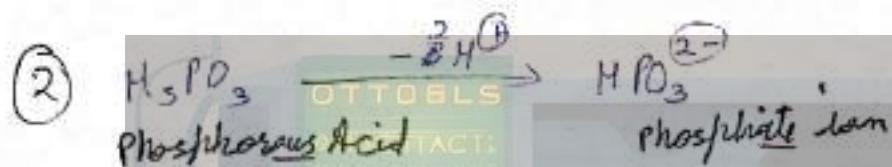
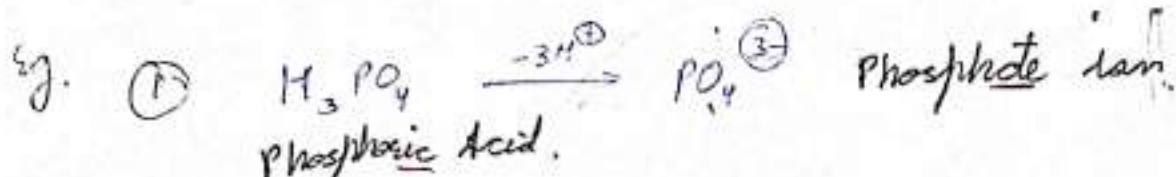
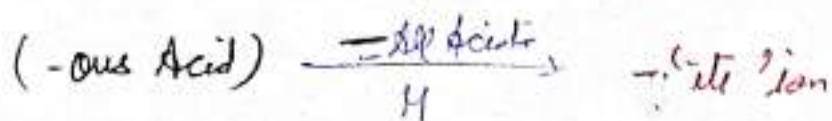
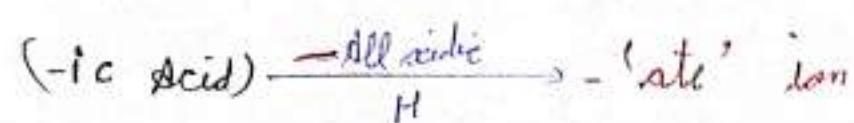


Basicity: - 0
O.S.: - 6

Not A Acid

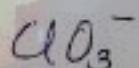
Naming Of Salts:-



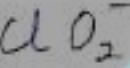


Q write formula

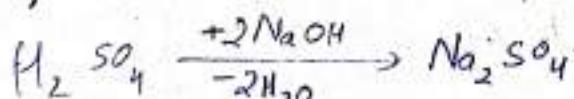
① Chlorate ion.



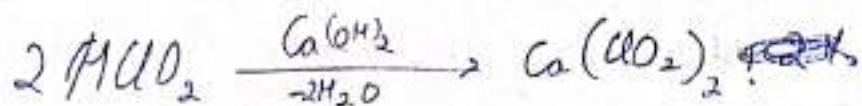
② Chlorite ion.



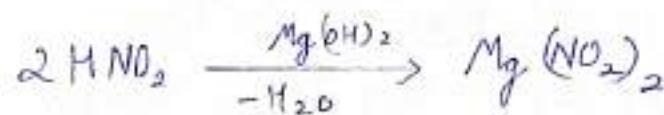
③ Sodium Sulfate



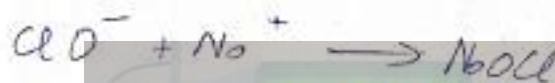
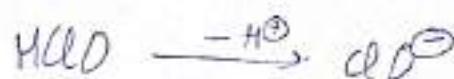
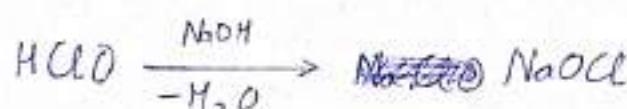
④ Calcium Chloride.



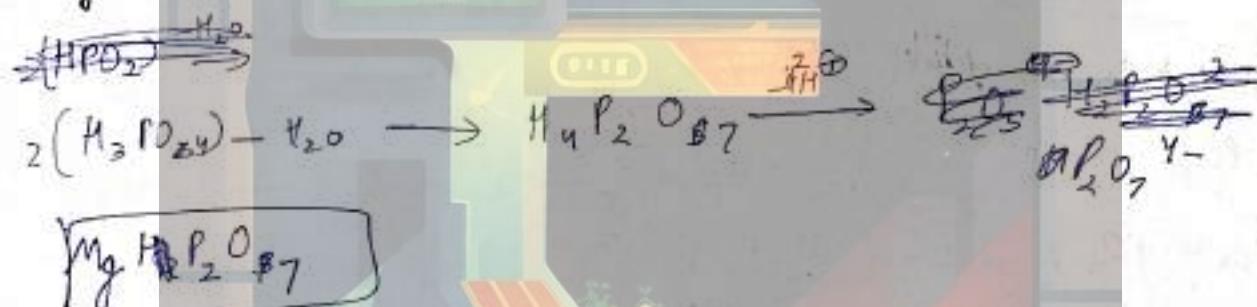
⑤ Magnesium Nitrite



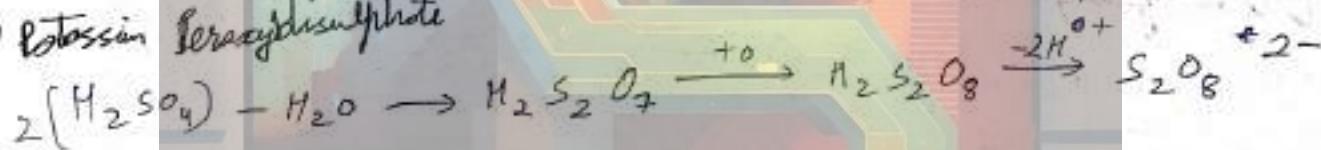
⑥ Sodium Hypochlorite



⑦ Magnesium Pyro Phosphate.



⑧ Potassium Persulfate

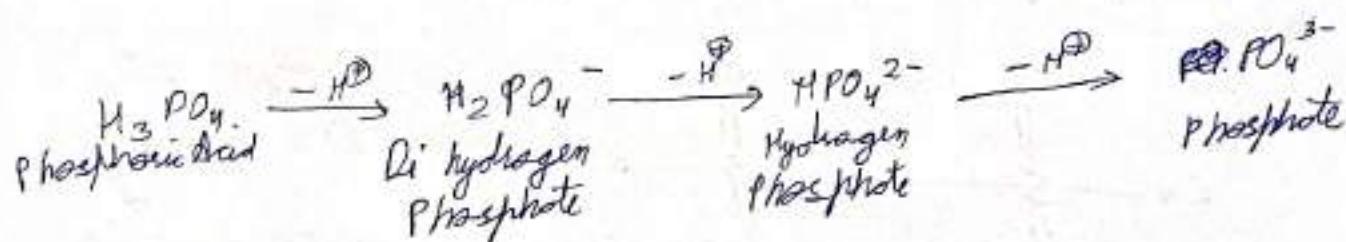


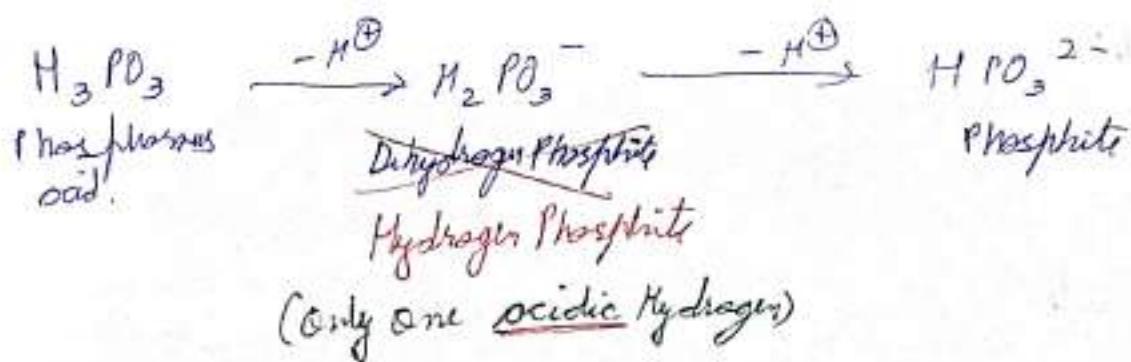
Naming of Salts with Acidic Hydrogen's

H_2SO_4
H₂S₂O₈

Hydrogen
Sulphate

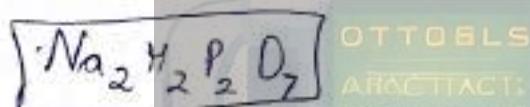
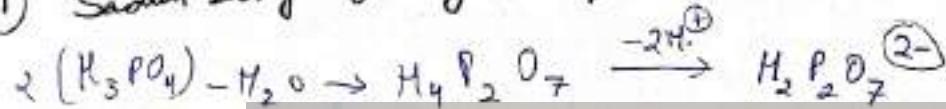
Hydrogen
Sulphate.



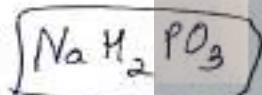
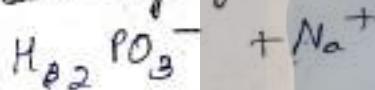


Q write Chemical formula

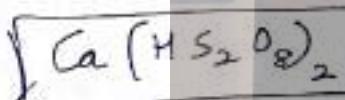
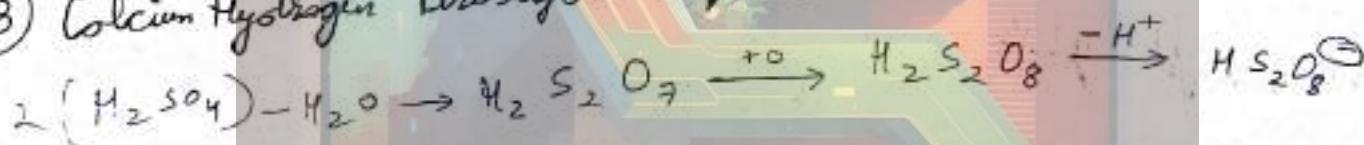
① Sodium Dihydrogen Pyrophosphate



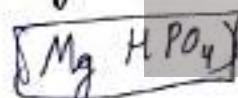
② Sodium Hydrogen Phosphate



③ Calcium Hydrogen Peroxydi-sulphate.



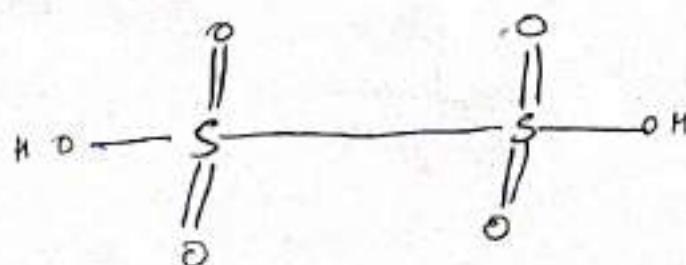
④ Magnesium Hydrogen Phosphate.



Thionic Acid \Rightarrow General Formula \rightarrow $\text{H}_2\text{S}_n\text{O}_c$ $n > 1$

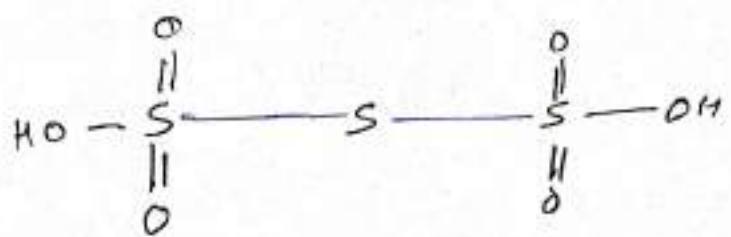
$$m = 2 \quad \text{H}_2\text{S}_2\text{O}_6$$

Di Thionic Acid



$$\begin{aligned} \text{S-S linkage} &= 1 \\ \text{average O.S} &= +\frac{10}{2} \\ \text{Basicity} &= 2 \end{aligned}$$

$n=3$ $H_2S_3O_6$ Tri-Sulfuric Acid



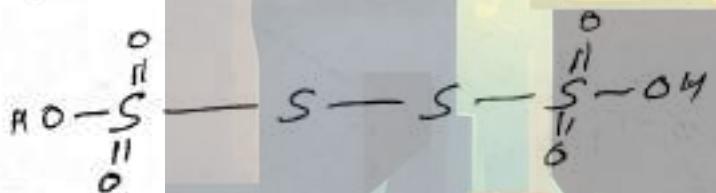
$$S-S \text{ linkage} = 2$$

$$\text{average } O-S = \frac{+10}{3}$$

$$\text{Basicity} = 2$$

$n=4$ $H_2S_4O_6$

OTTOBLS
ARA
Tetra Thionic acid



$$S-S \text{ linkage} = 3$$

$$\text{average } O-S = \frac{+10}{4}$$

$$\text{Basicity} = 2$$

General:-

$$S-S \text{ linkage} = n-1$$

$$\text{Average } O-S = \frac{+10}{n}$$

$$\text{Absolute.} = +5, 0 \dots n, +5$$

$$\text{Basicity} = 2$$

Q Write chemical formula of Sodium Hydrogen Dithionate

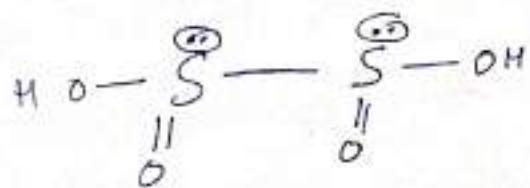
$\boxed{\text{NaHS}_2\text{O}_6}$

Thismous Acid

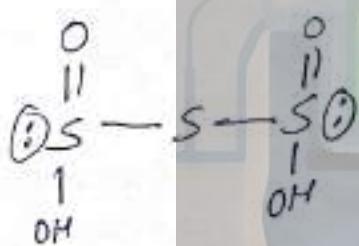
General: $H_2S_{n-1}O_n$ ($n > 1$)

$n=2$ Dithiomous Acid.

: O.S: - +3, +3



$n=3$ Triethiomous Acid.



OTTOBL'S
AROMATIC
CO. LTD.
MANUFACTURERS

O.S: - +3, 0, +3

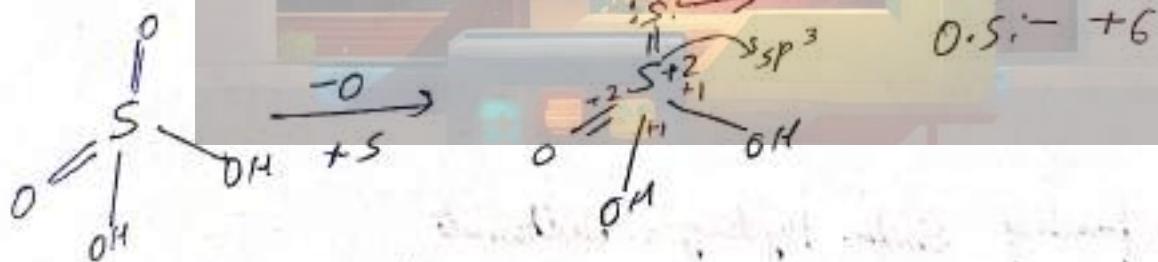
No. of S-slinks :- $n-1$.

Avg O.S

Absolute

Basicity :- 2

This-sulfurous acid. ($H_2S_2O_3$)

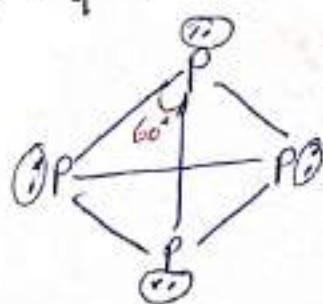


① Sodium This Sulphate

$Na_2S_2O_3$

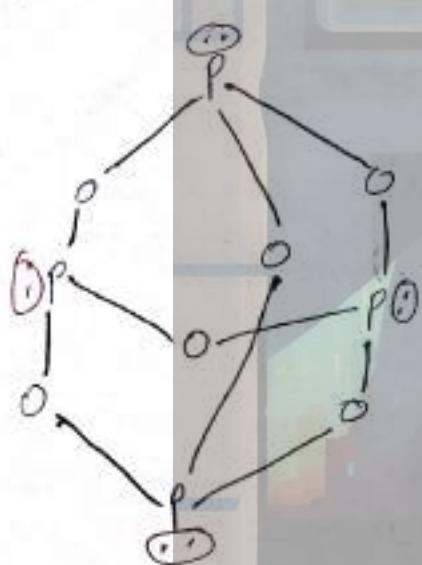
Important Structures:-

① P_4 (white Phosphorus)



p-p bonds $\rightarrow 6$
Bond Angle $\rightarrow 60^\circ (12)$

② $P_4 O_6$

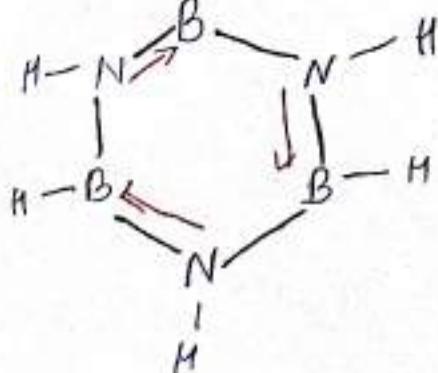
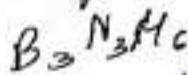


③ $P_4 O_{10}$

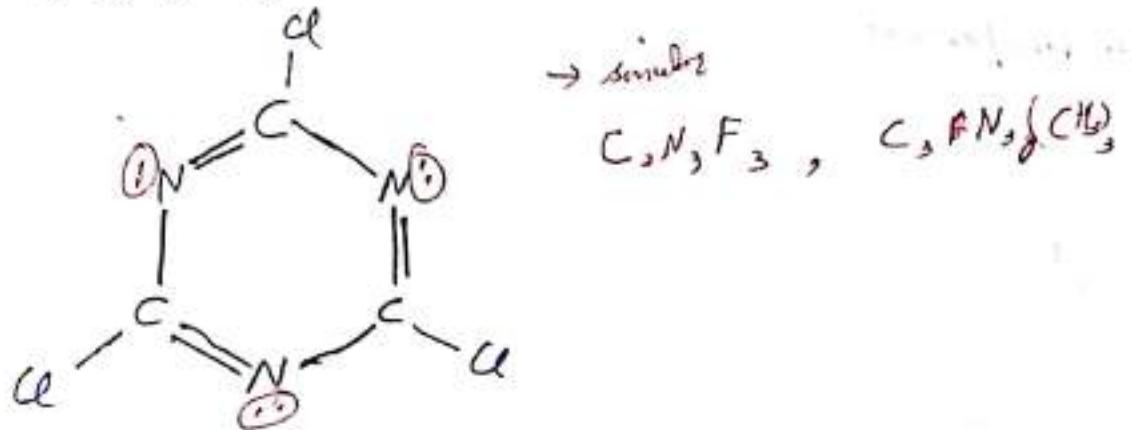


$P_4 O_7, P_4 O_8, P_4 O_7$ forms in similar ways.

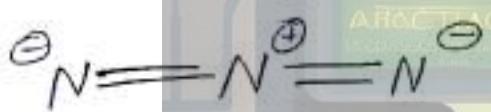
④ Inorganic Benzene & Borazine



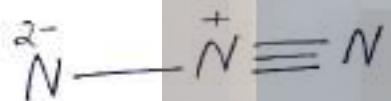
⑤ $C_3N_3O_3$ (Cyanic Trichloride)



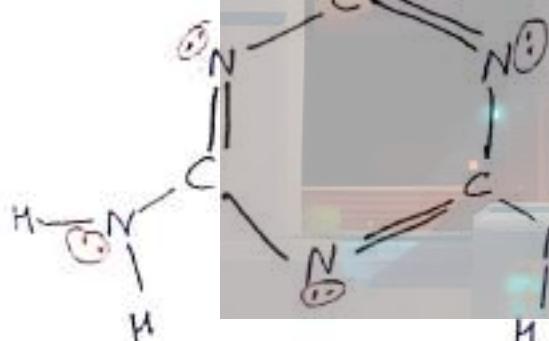
⑥ N_3^- (Azide)



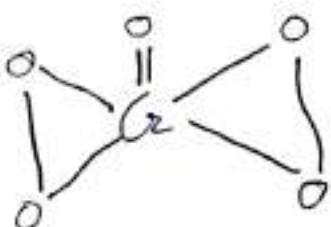
O_2

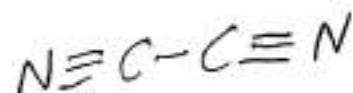
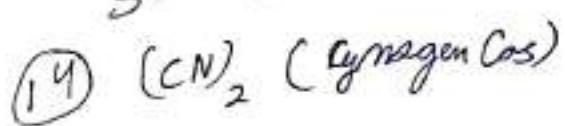
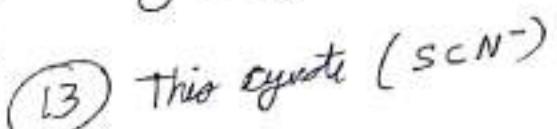
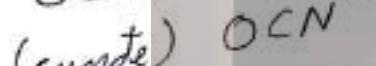
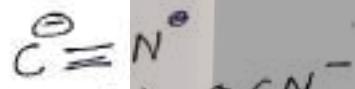
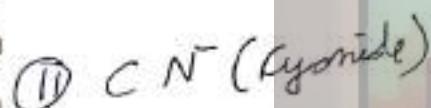
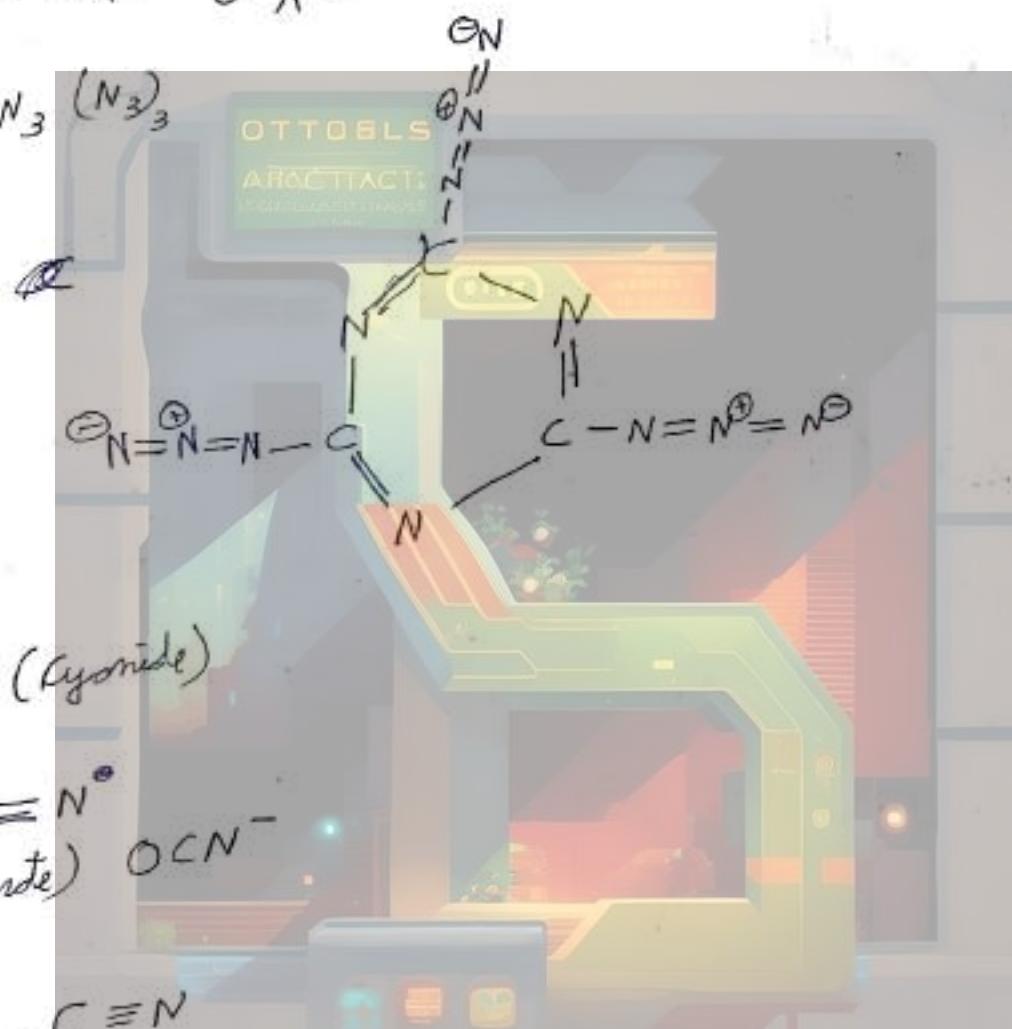
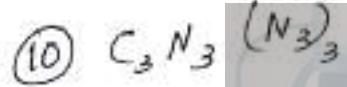
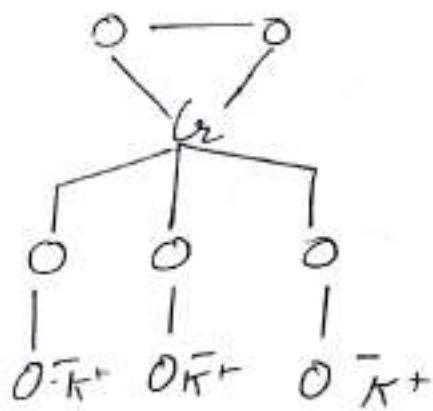
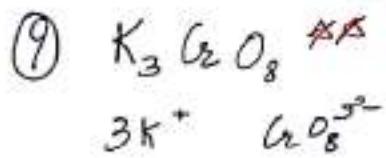


⑦ $C_3N_3(NH_2)_3$ (Lysine Triamine) (Melamine)

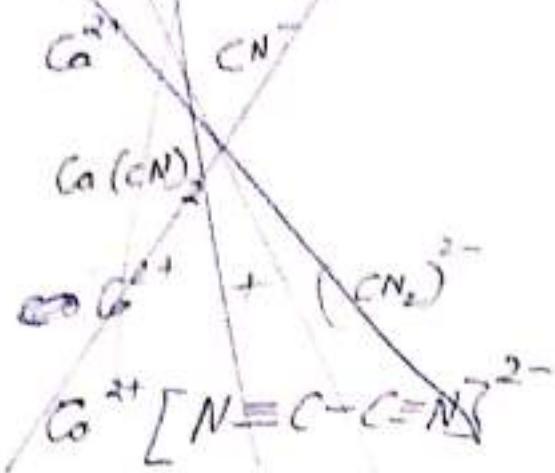


⑧ CrO_5

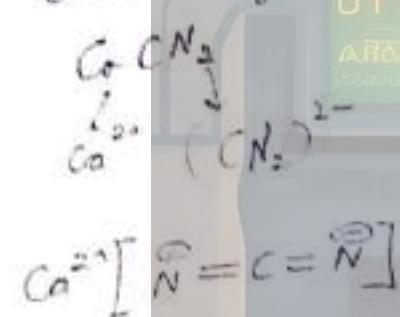




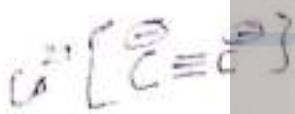
15) Calcium cyanamide



15) Calcium cyanamide



16) Calcium carbide (CaC_2)



17) Tetracyan Ethylene $[\text{C}_2(\text{CN})_4]$

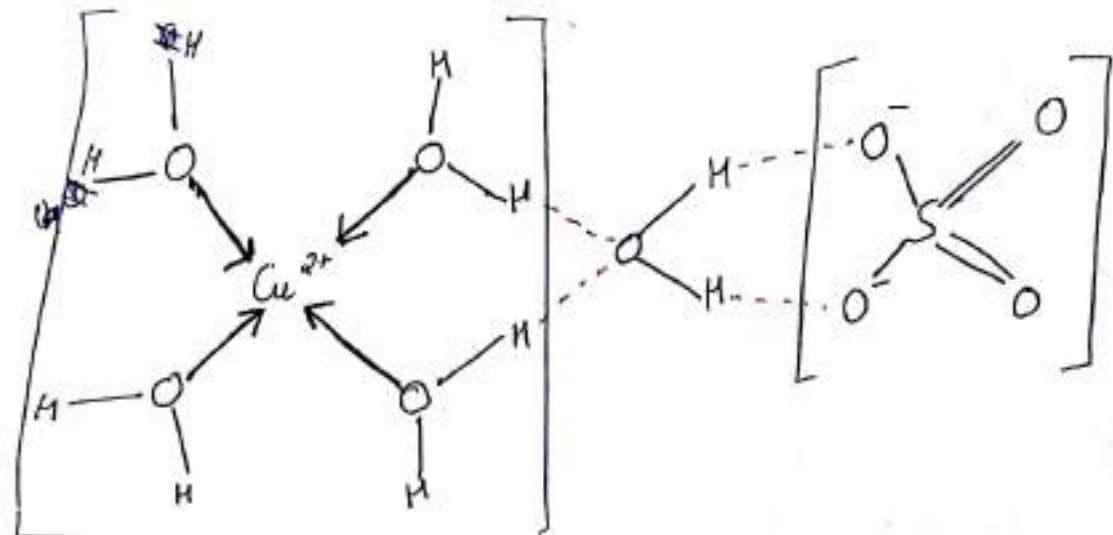


18) HN_3 (Hydrazoic acid)



170)

⑯ Blue Vitrol ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$)



Types of Water Molecules → 3

coordinating, covalent — 2

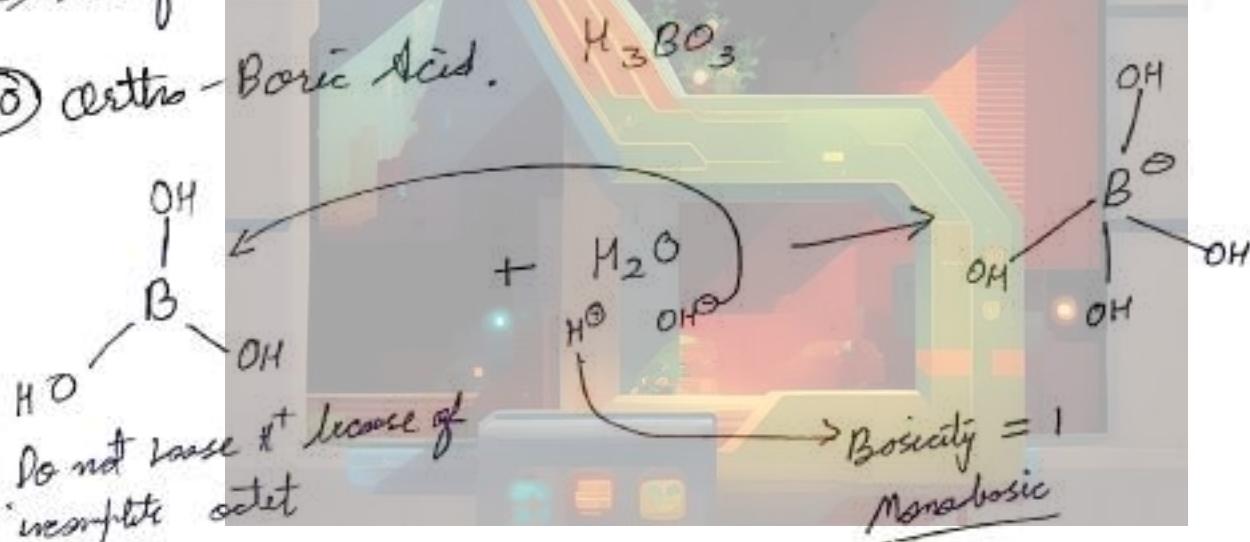
Coordinating, covalent & Hydrogen — 2

Hydrogen & covalent — 1

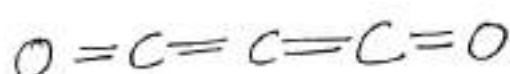
→ Ionic, covalent, co-ordinate, H-bond present in $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

→ No. of H_2O molecules off crystallization = 5

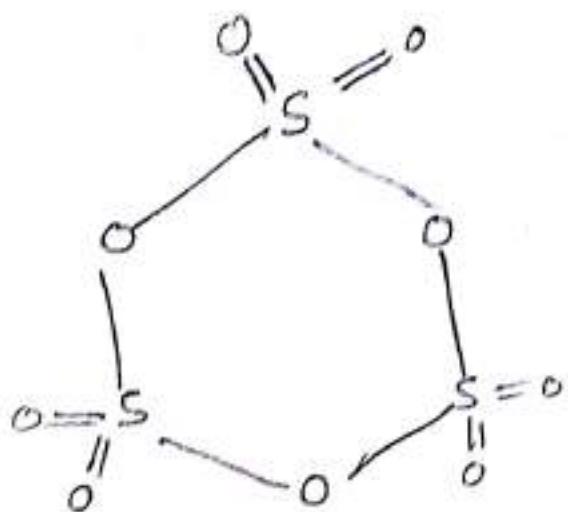
⑰ Ortho-Boric Acid.



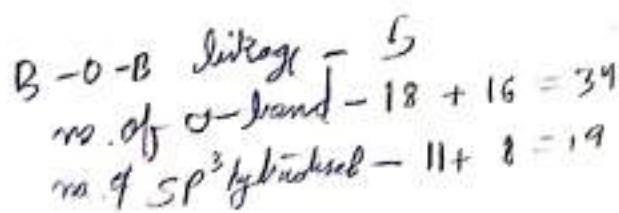
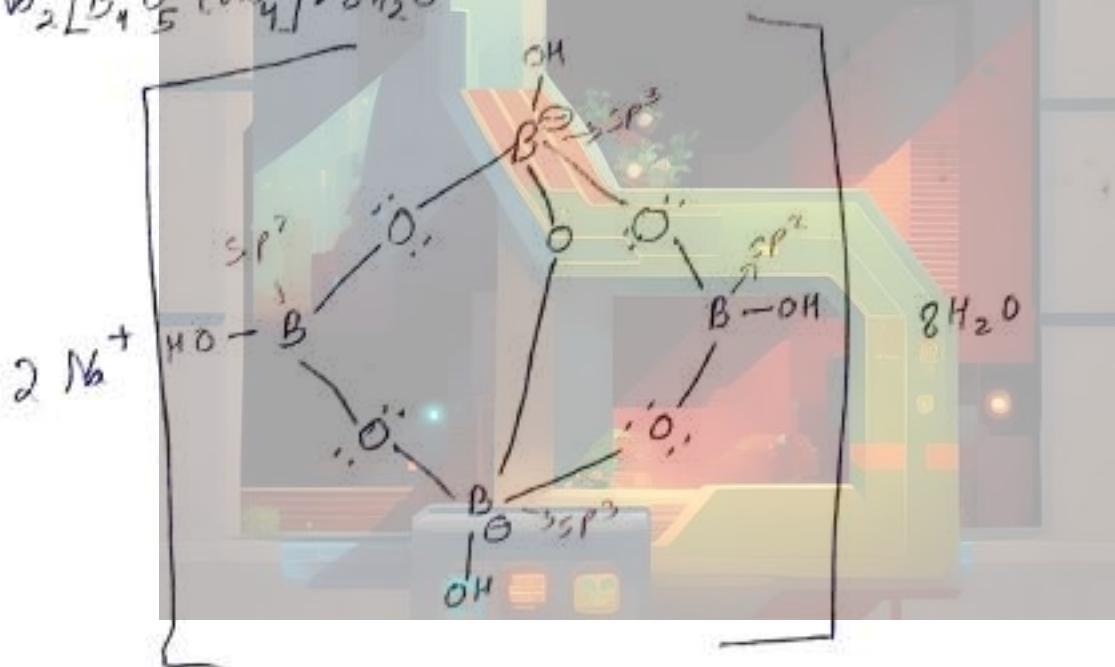
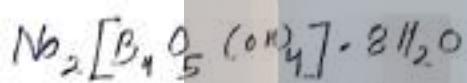
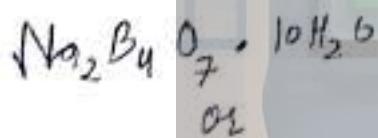
⑱ Carbon Sub oxide. C_3O_2



(22) S_8

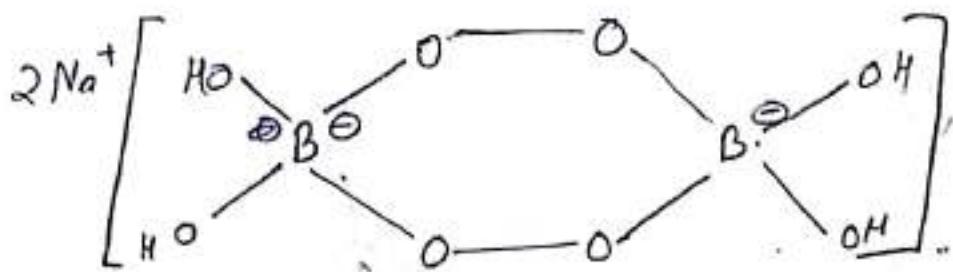


(23) Borox eH



(172)

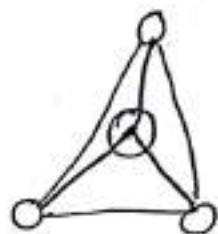
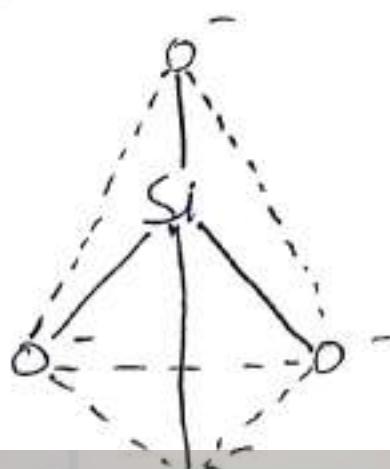
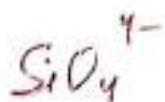
(24) Sodium Peroxo Borate $\text{Na}_2\text{B}_2\text{O}_4(\text{OH})_4$



Silicates

- Silicate is made up with silicon and oxygen atoms
- In silicate every silicon atom is surrounded by 4 oxygen atoms which are arranged at corner of tetrahedron & bonded with single bond.
- There are 6 types of silicates & all are non planar.
 - ① Cyclic/Sesqui Silicate
 - ② Pyro/Soro Silicate
 - ③ Cyclic Silicate
 - ④ Chain Silicate
 - a) Single Chain Silicate [Lyonite Silicate]
 - b) Double Chain Silicate [Amphibole Silicate]
 - ⑤ 2-D Silicate or Sheet Silicate [Phyllo Silicate]
 - ⑥ 3-D Silicate [Tecto-Silicates]

① Ortho Silicate / Neo Silicate

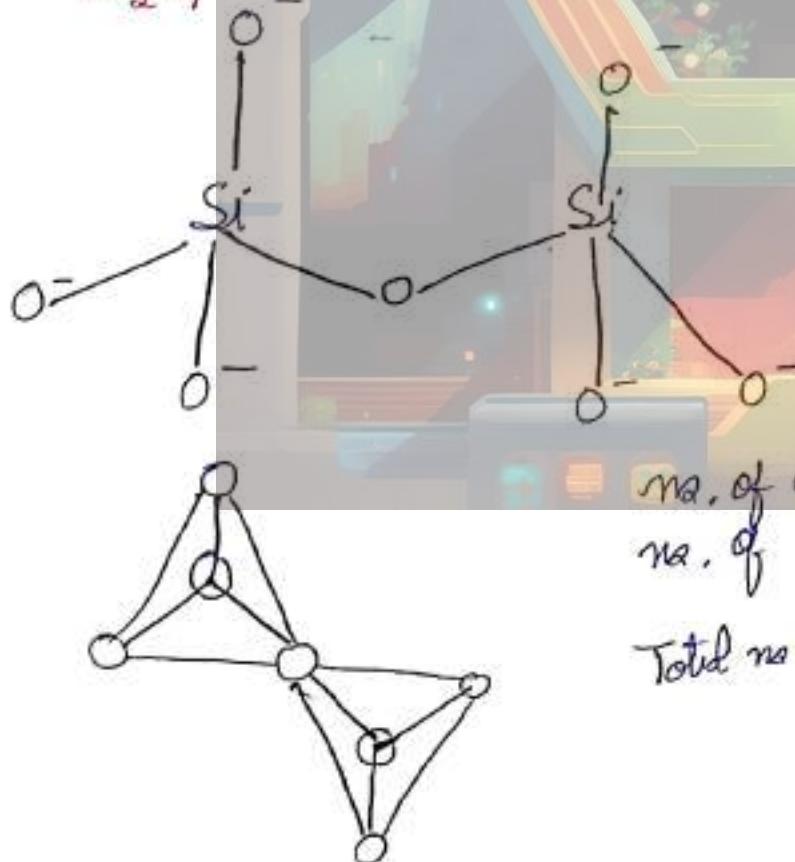
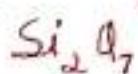


• → silica

O → oxygen

no. of \equiv oxygen per tetrahedron :- 4
no. of \equiv oxygen contribution per tetrahedron :- 4

② Pyro Silicate / Soro Silicate



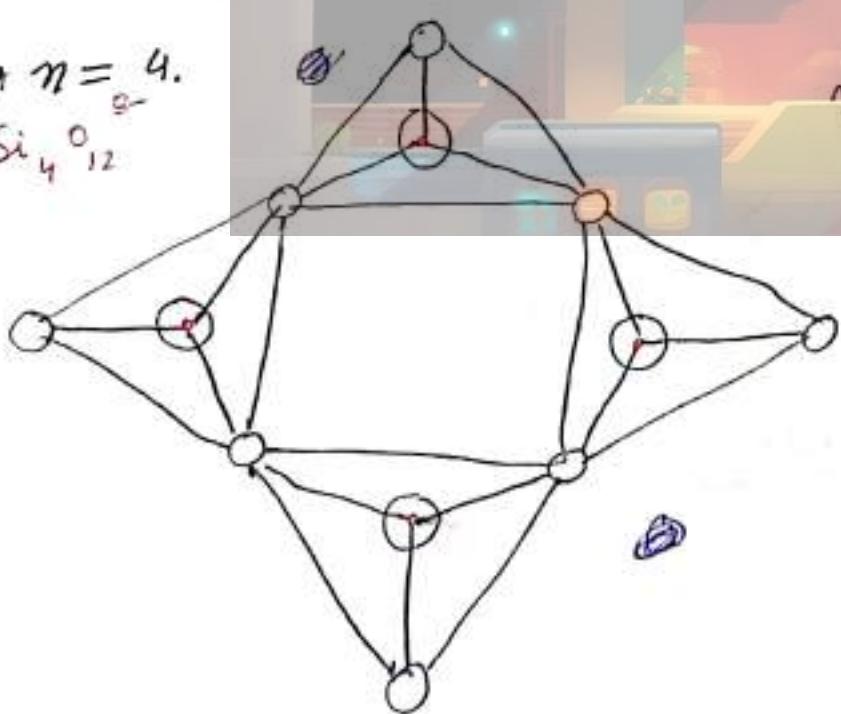
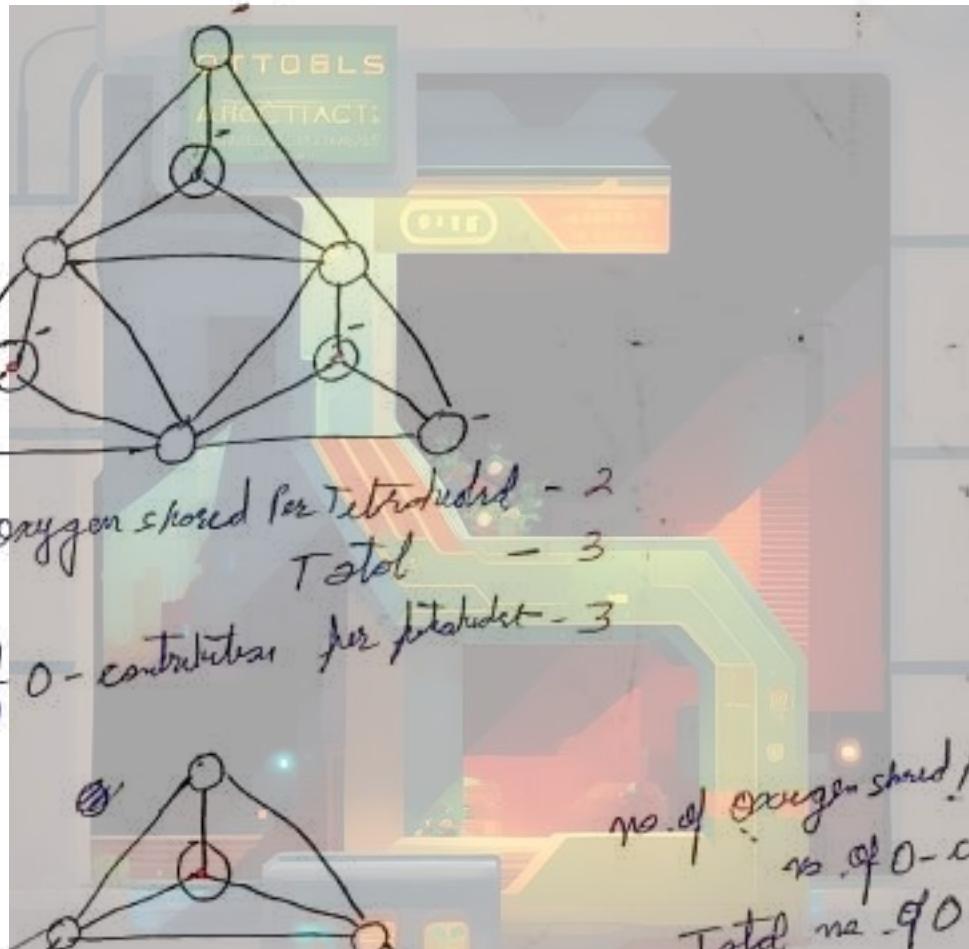
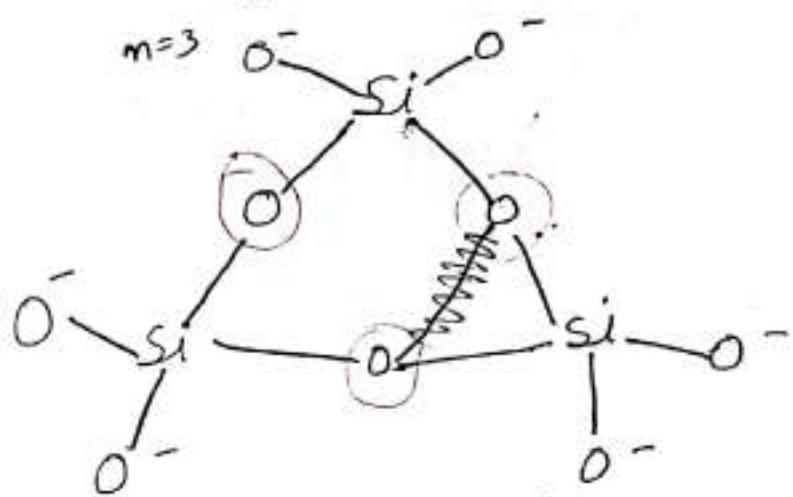
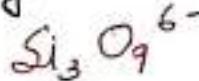
shared

no. of oxygen \equiv per Tetrahedron - 1

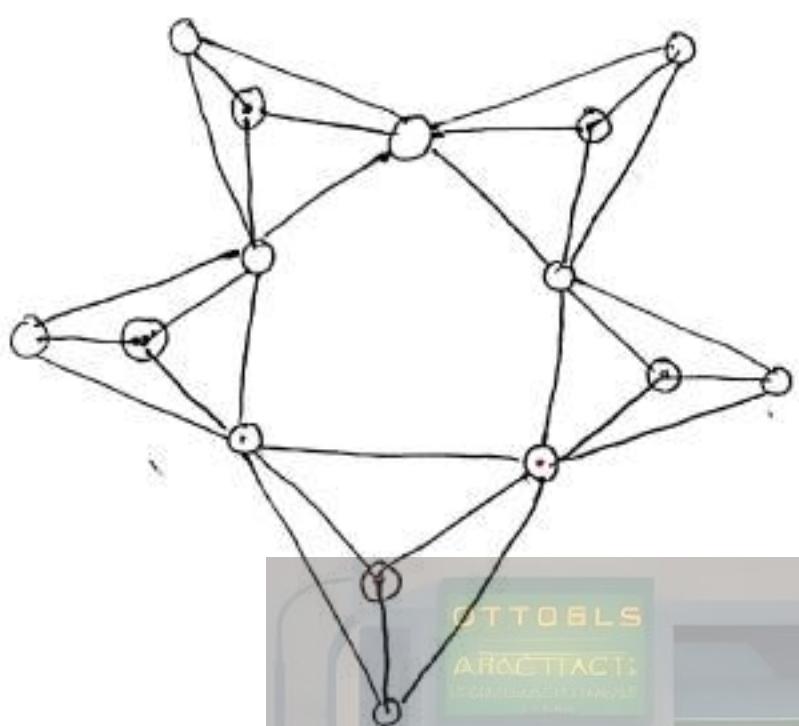
no. of oxygen contribution Per Tetrahedron - 3 - 5

Total no. of oxygen shared - 1

③ Cyclic Silicide / Ring Silicate



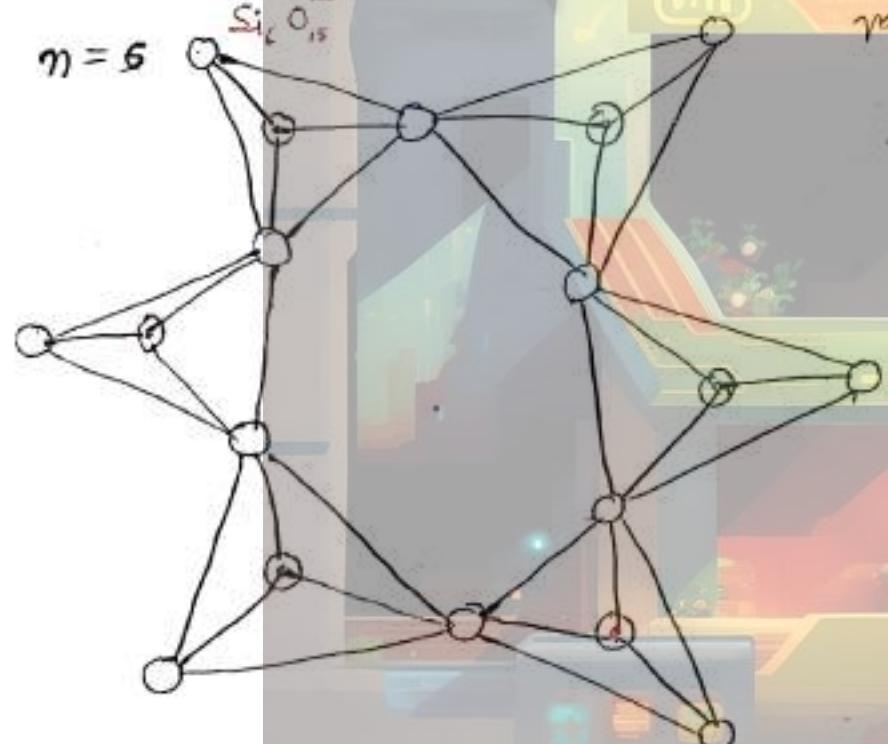
$m=5$ $\text{Si}_5\text{O}_{15}^{10-}$



Shared
no. of oxygen per tetrahedron = 2
Total oxygen shared = 5
Oxygen centre per tetrahedron = 3

$n=6$

$\text{Si}_6\text{O}_{15}^{12-}$



no. of oxygen shared per tetrahedron = 2
Total oxygen shared = 6
Oxygen centre per tetrahedron = 3

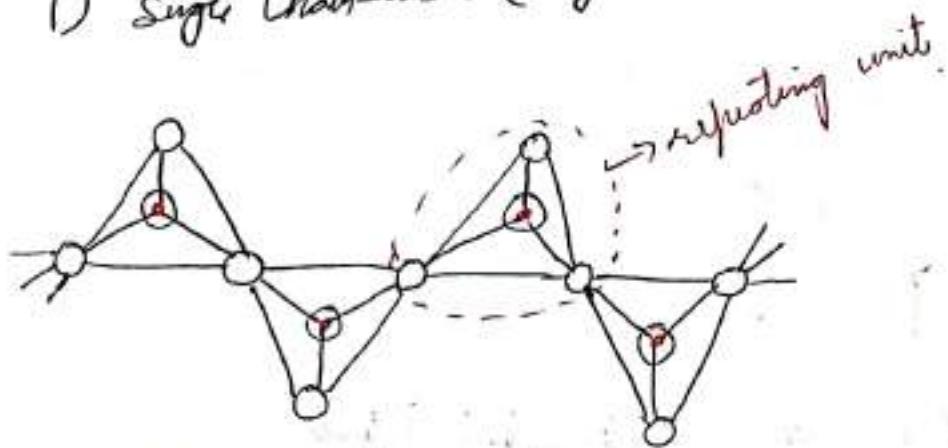
General :-

$\text{Si}_n\text{O}_{3n}^{-2n}$

no. of oxygen shared per tetrahedron = 2
no. of oxygen contribution per tetrahedron = 3
Total no. of oxygen - shared = n

④ Chain Silicates

i) Single Chain Silicate (Pyroxene Silicate)



formula for one unit - $\text{Si}_4\text{O}_{10}^{2-}$

for one chain -
AHOCTACTI

General Formula:-

$$(\text{Si}_4\text{O}_{10}^{2-})_n \rightarrow \text{Si}_n\text{O}_{3n}^{-2m}$$

silicate

no. of oxygen shared per tetrahedron :- 2

no. of oxygen shared per centre per tetrahedron :- 3

Note - If single chain silicate is restricted from both ends.

$$(\text{Si}_n\text{O}_{3n}^{-2m}) + \text{O}_2^-$$

$$\text{Si}_{n+1}\text{O}_{3n+1}^{-2m-2}$$

is terminal chain.

Q ① True or False

Formula of cyclic silicate is same as single chain silicate

Silicates are same. True

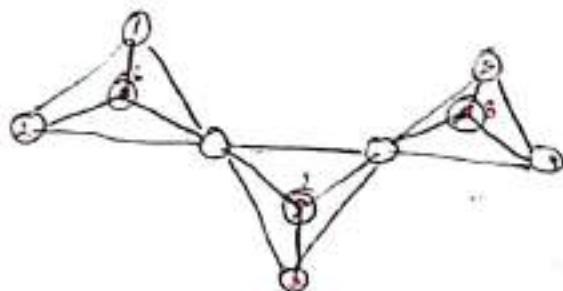
② Total no. of oxygen shared in 3 membered cyclic silicate & 4 membered single chain silicate are not same. ~~True~~ False

③ Formula of 4 member cyclic silicate & 4 member single chain silicate are same. ~~False~~ True

$\text{Si}_4\text{O}_{12}^{-8}$
cyclic

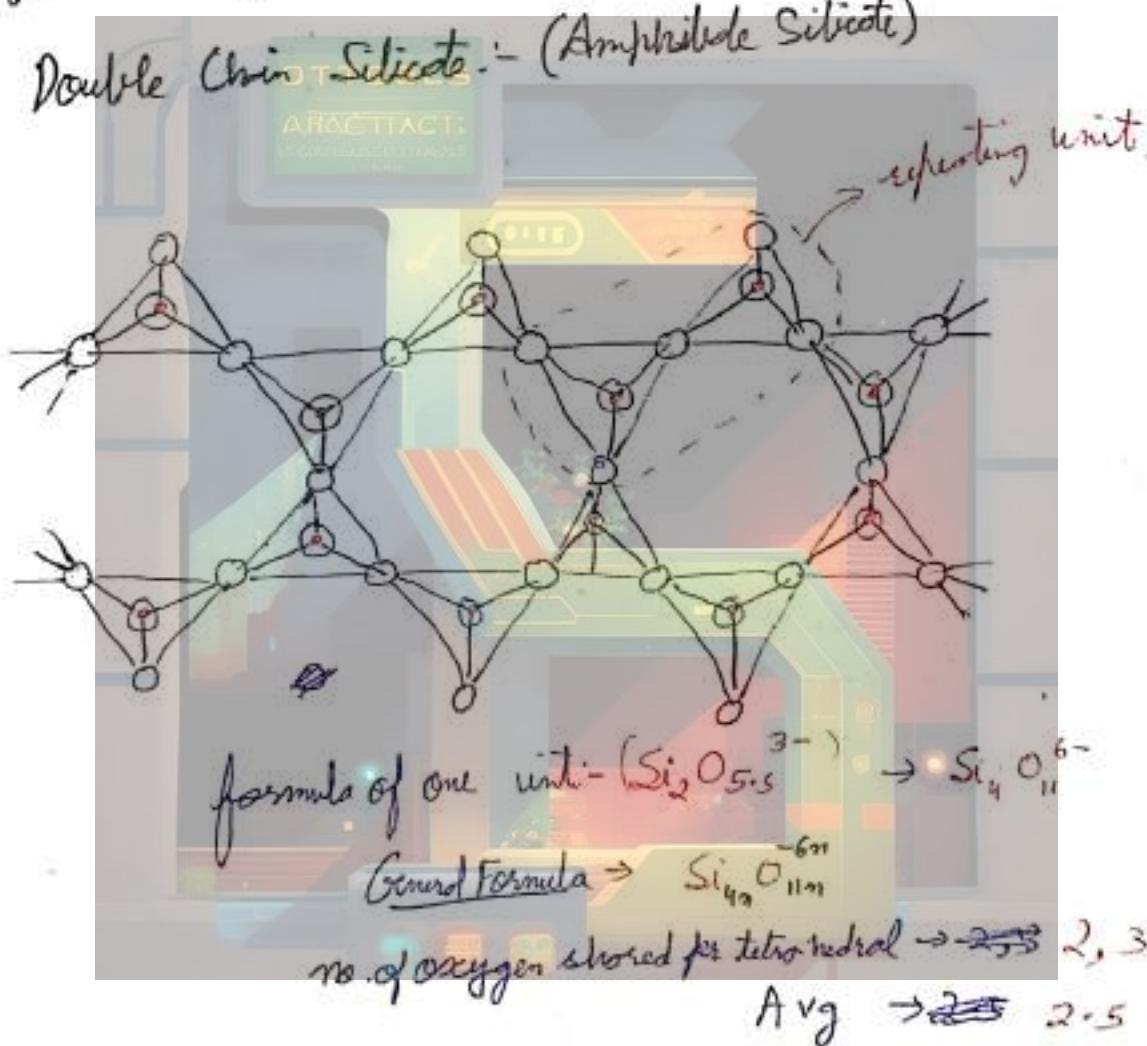
$\text{Si}_4\text{O}_{13}^{-10}$
4-membered single chain silicate

④ Total no. of unbonded oxygen in 3 member single chain silicate
or 7. False.

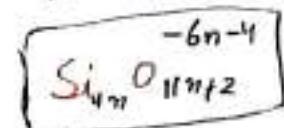
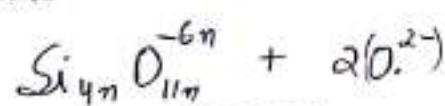


Q find formula of 6 member single chain Silicate. $\text{Si}_6\text{O}_{14}^{-14}$

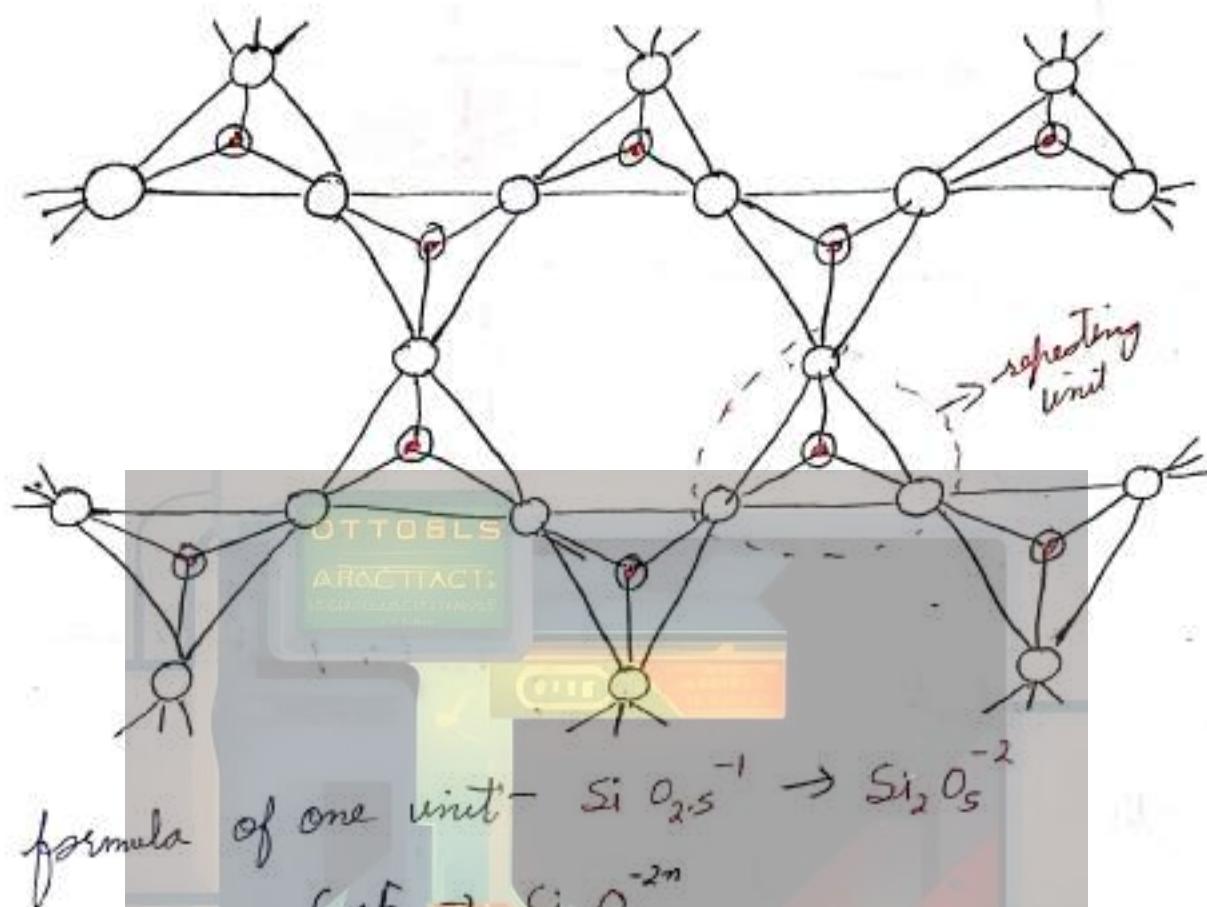
ii) Double Chain Silicate :- (Amphibole Silicate)



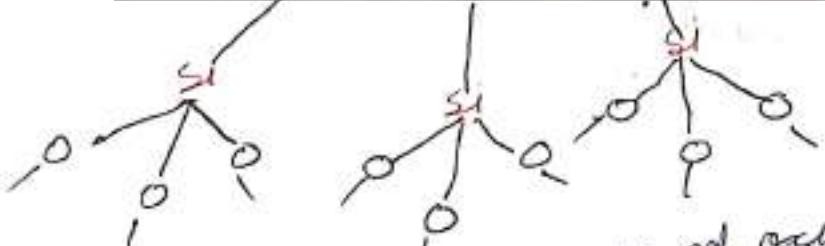
Note:- If double chain silicate is restricted from both ends.



⑤ Sheet Silicate / 2D-Silicate / Phyllosilicate

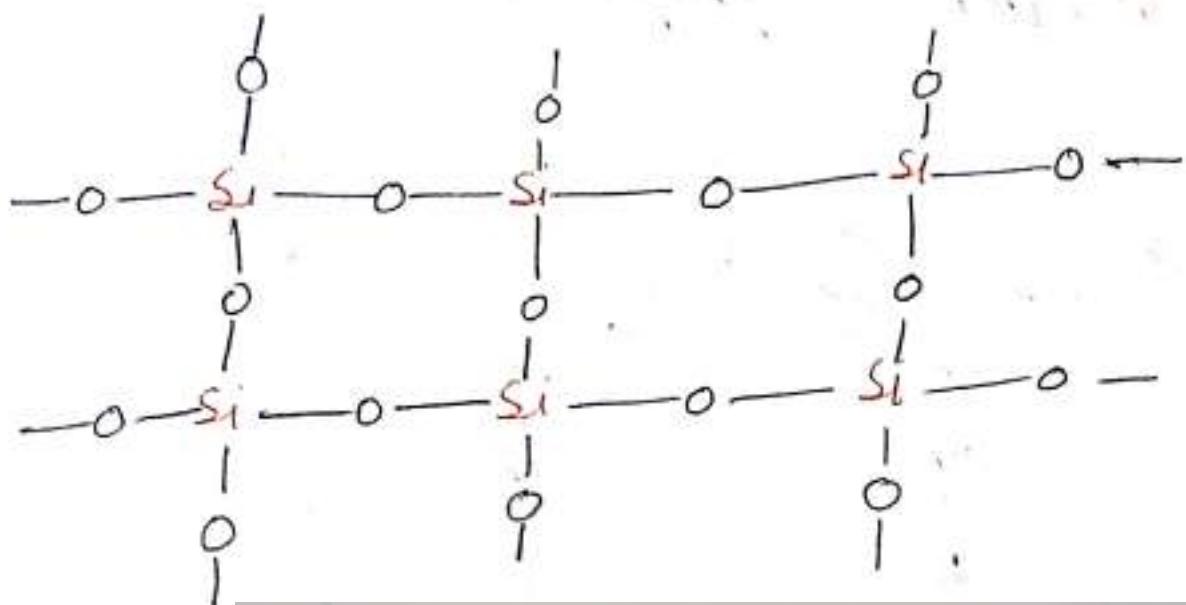


⑥ 3D-Silicate (Tetrahedral Silicate)



G.F. - $(\text{Si O}_2)_n$

no. of oxygen shared per Tetrahedron - 4
No. of O-centres per tetrahedron - 2



Silica (SiO_2)

OTTOBL'S
ROCKS &
MINERALS

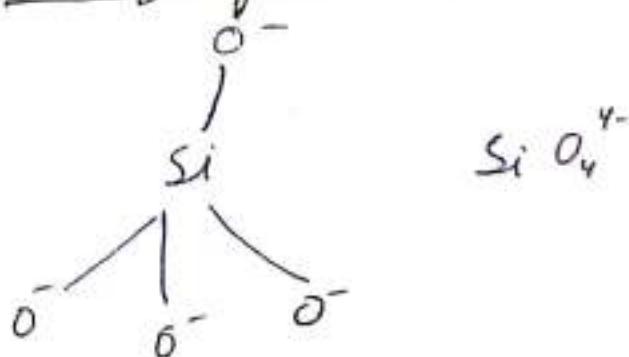
Silicates

No. of shared O per tetrahedron

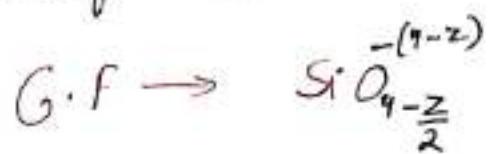
1. Ortho
2. Pyro
3. Cyclic
4. Single Chain
5. Double chain
6. 2-D
7. 3-D



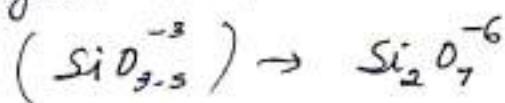
Generalized formula of Silicates :-



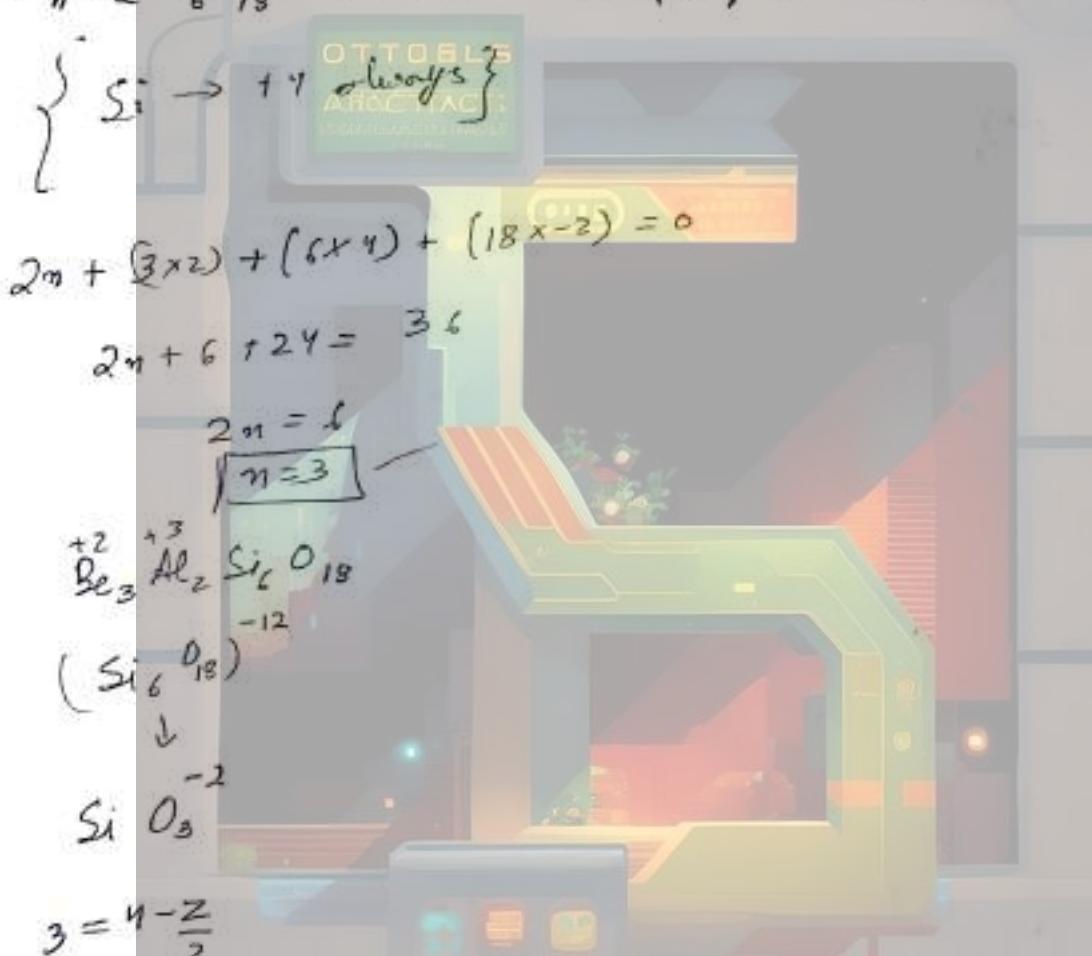
Z no. of O-stained Polyhedron,



Ex. Lyso. $\rightarrow Z=1$



Q $Be_3Al_2Si_6O_{18}$ Calculate 'n' & Type of silicate.



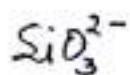
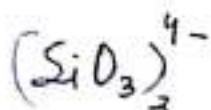
$$\frac{Z}{2} = 1$$

$$Z = 2$$

Cyclic / Single Chain

§ Identify Type of Silicate

① $\text{LiAl}(\text{SiO}_3)_2$



$$\cancel{2\text{O}} - 2 = z - 4$$

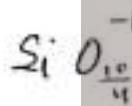
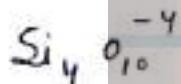
$$z = 2$$

Cyclic / Single Chain

② $\text{Mg}_3\text{Si}_4\text{O}_{10}(\text{OH})_2$

$$6\text{O} - 2 + x = 0$$

$$x = -4$$

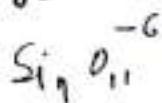


$$-1 = z - 4$$

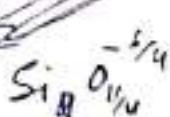
$$\boxed{z = 3}$$

2D / sheet

③ $\text{Ca}_2\text{Mg}_5(\text{Si}_4\text{O}_{11})_2 \cdot 6\text{H}_2\text{O}$



$$-6 = z - 4$$



$$\cancel{2} - \frac{3}{2} = z - 4$$

$$z = \frac{8-3}{2}$$

$$z = \frac{5}{2}$$

Double Chain

④ SiO_2 (Silico)

3-0

⑤ $2\text{e} \text{ SiO}_4$ (charge of 2e)

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

$Z=0$ ortho

$$\text{charge} = Z - 4 = -4$$

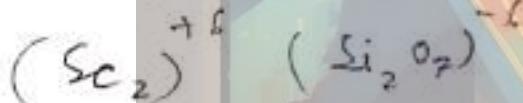
⑥ $\text{Sc}_2 \text{Si}_2 \text{O}_7$

$$11 - \frac{Z}{2} = \frac{7}{2}$$

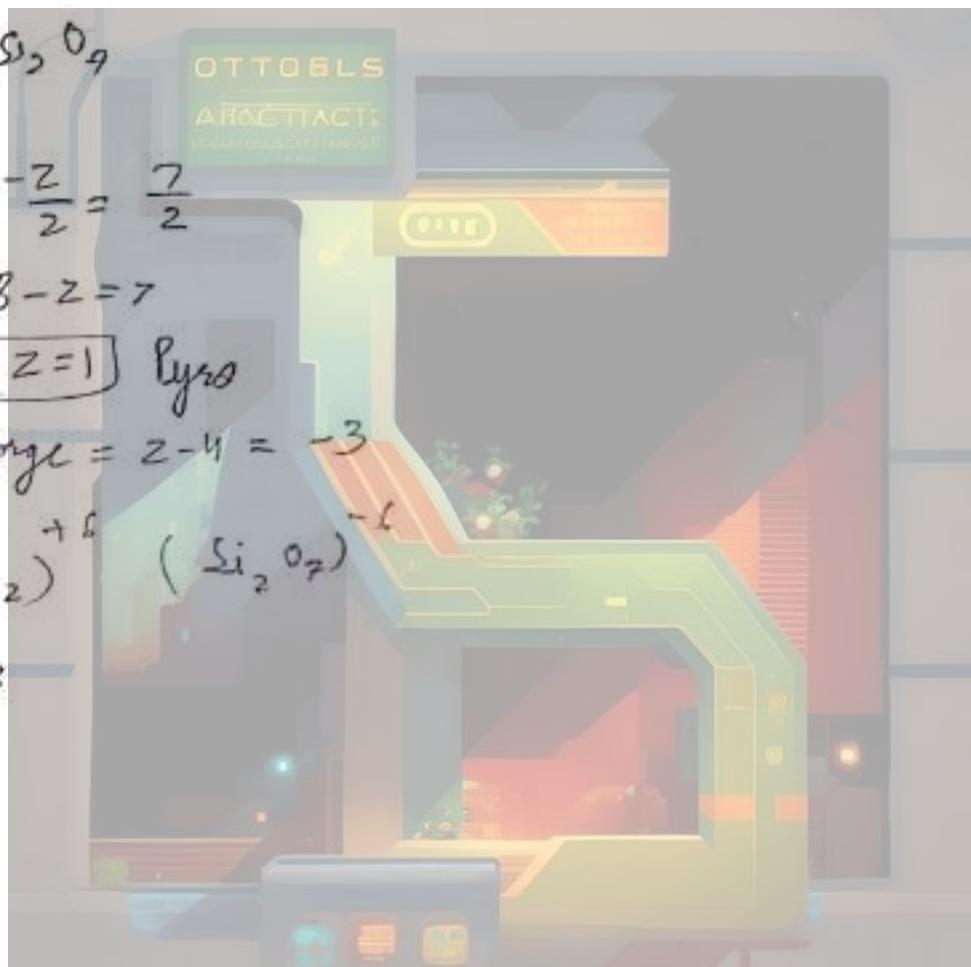
$$8 - Z = 7$$

$Z=1$ Pyro

$$\text{charge} = Z - 4 = -3$$



$$\text{Sc} \rightarrow +3$$



Allotropes (read from module)

→ Existence of a chemical element in 2 or more other forms.

Allotropes of Carbon

- ① Diamond
- ② Graphite
- ③ Fullerene

crystalline allotropes.

follow regular geometry.

① Diamond

→ Crystalline Structure

ABCTACTI

→ 3D-structure

→ Each carbon atom attached with other 4 carbon atoms

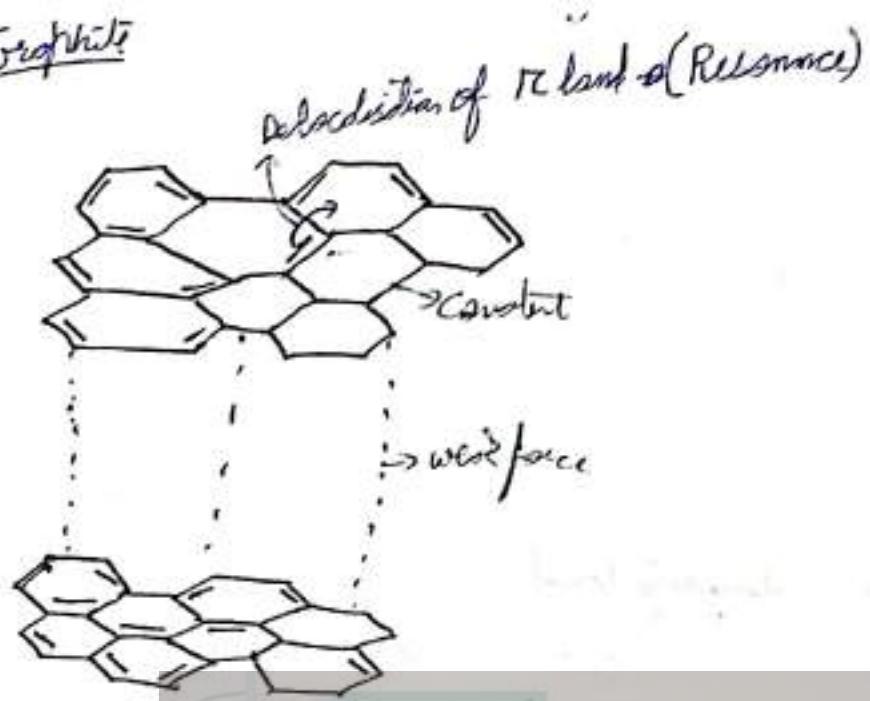
like a tetrahedral unit hence each carbon is sp^3 hybridised.



→ Naturally Hardest substance

→ C-C bond length 154 pm

② Graphite



- Layered Structure
- Each carbon is attached with other carbon hence Hybridisation of each carbon is sp^2 .
- The remaining $4^{th} e^-$ can form π -bands
- Graphite is lustrous (shiny & soft)
- Electrons are de-localised along the sheet hence graphite is good conductor of electricity.
- It is also used as dry lubricants

③ Fullerene

↓
purest form of carbon; due to absence of "Dangling Bond".

- Cages / Football like structure
- Consists of 5-membered + 6-membered rings.

★ C₆₀ (Buckminsterfullerene)

→ 12 5-member rings, 20 6-member rings

→ Each carbon is sp^2 hybridised no. of vertices = 60

Dangling Bond - Unsatisfied valencies of element in network solid.

O-1 $\{39 \text{ to } 48, 50\}$ O-2 $[21, \cancel{2} 36]$ S-1 $\{10, 11\} + \{14, 15\}$ S-2 $\{14, 16\}$ J.M $\{9 \text{ to } 13, 15, 17 \text{ to } 20, 32, 34 \text{ to } 37, 39 \text{ to } 41, 49, 50, 59\}$

Q) which does not have dangling bond.

- a) SiC b) Diamond c) SiO_2

✓ d) fullerene

Diamond like structure (lattice structure)

In General:- C_n [Fullerene]
 no. of pentagonal rings = 12

$$\text{no. of hexagonal rings} = \frac{n}{2} - 10$$

Thermodynamic Stability
 Graphite > Diamond > Fullerene (c_{60})

Kinetic Stability

Diamond > Graphite

Bond Length

Diamond > Graphite



Allotropes of Oxygen

O_2
(oxygen)

O_3 (ozone)

Oxygen

- Colourless in gas phase
- Blue obs in liquid phase
- Paramagnetic nature

Ozone

- Colourless
- Diamagnetic nature

Allotropes of Sulfur

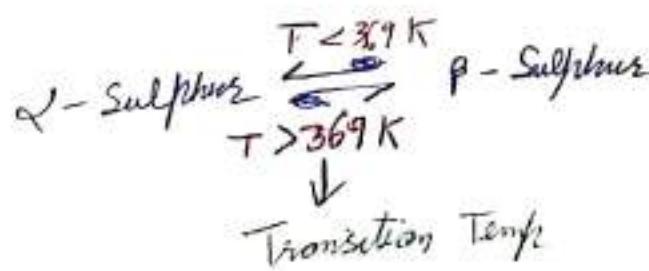
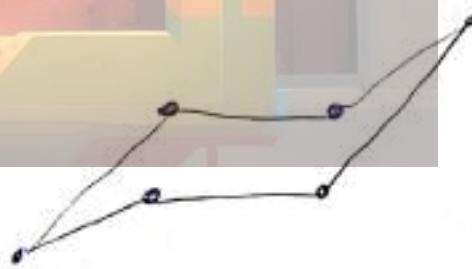
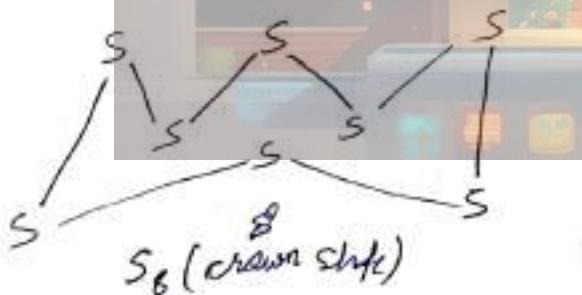
Rhombo. Sulphur
(α -Sulphur)

Yellow

Monoclinic Sulphur
(β -Sulphur)
Colourless

Thermodynamic Stability
Rhombo. Sulphur > Monoclinic Sulphur

→ Both α & β Sulphur have S_8 and S_6 form



Allotropes of Phosphorus

White
Phosphorus

Red
Phosphorus

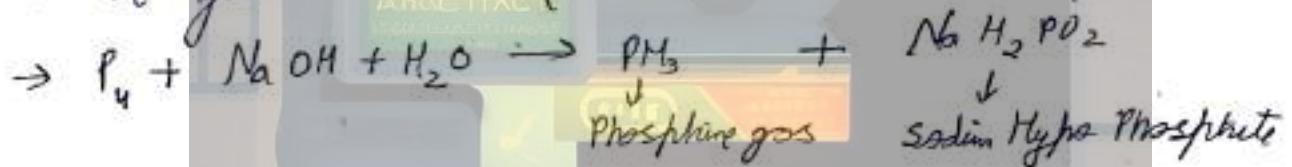
Black
Phosphorus

① white Phosphorus.

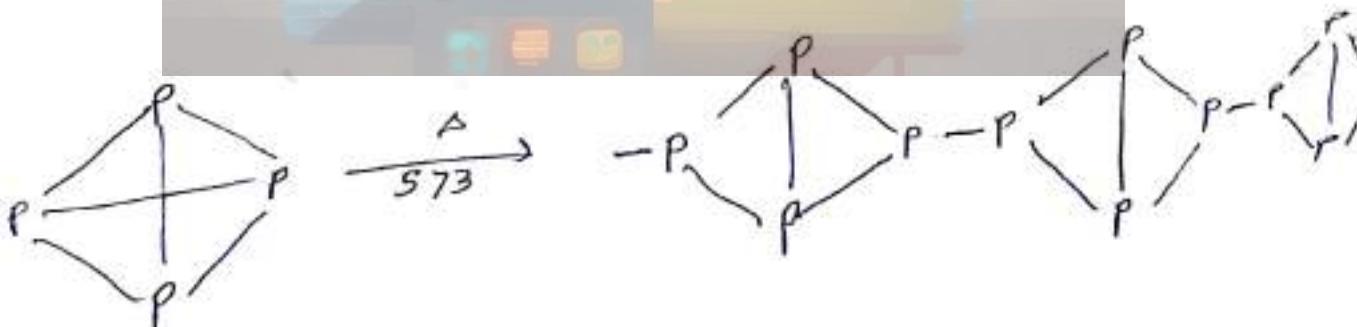
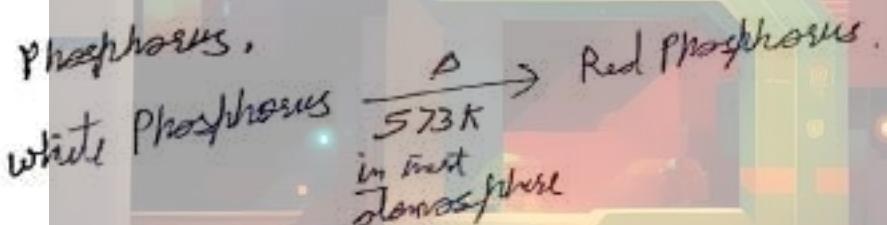
→ White translucent, waxy solid, poisonous

→ Soluble in C_6H_6 insoluble in H_2O

↔ It glows in dark $\xrightarrow[\text{Aerobic attack}]{\text{O}_2}$ (Chromiluminescence).



② Red Phosphorus.



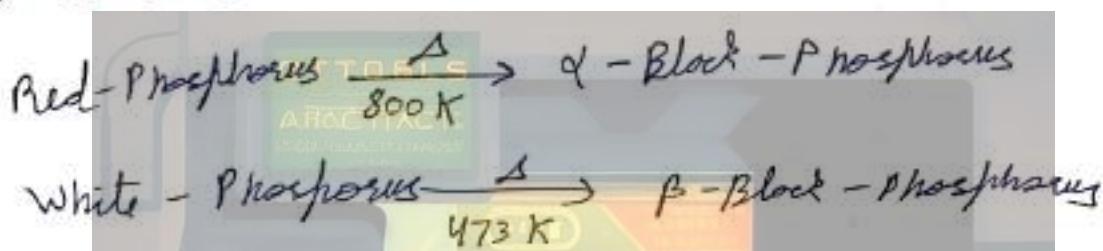
- non-poisonous, P_2
- Insoluble in H_2O as well as $\text{C}_5\text{H}_5\text{N}$
- Does not glow in dark.

Black Phosphorus

γ -Block Phosphorus

β -Block Phosphorus

→ 3D-structure



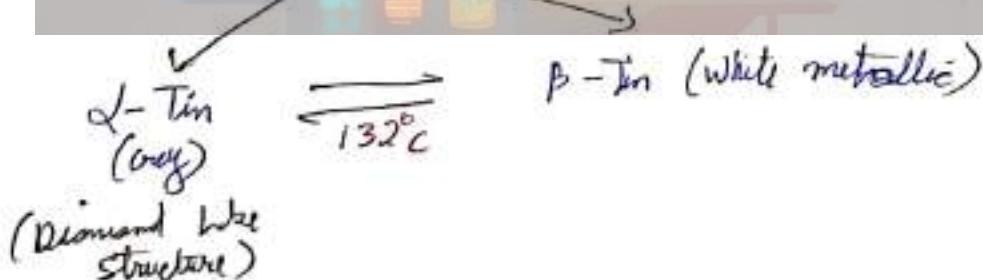
Thermodynamic Stability

White Block > Red Block > White

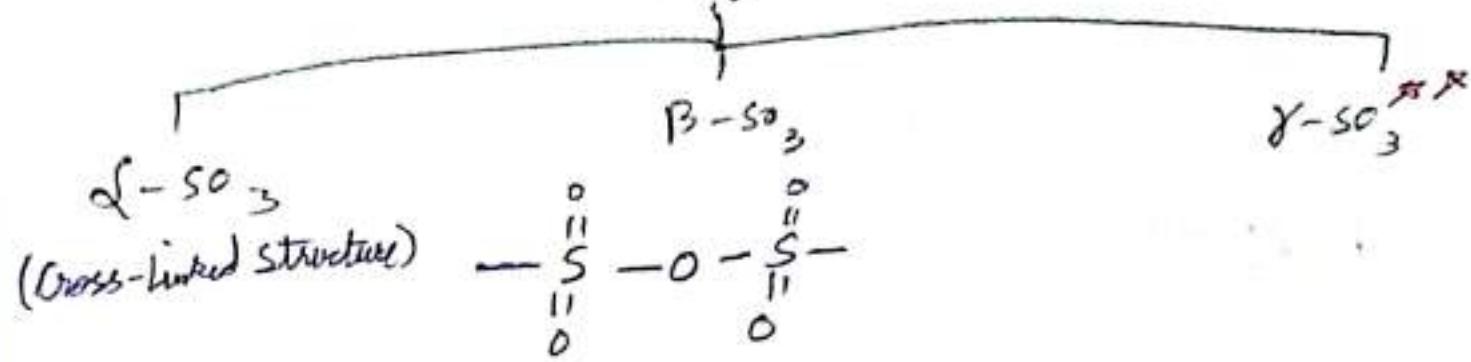
Reactivity Order

white > Red > Block

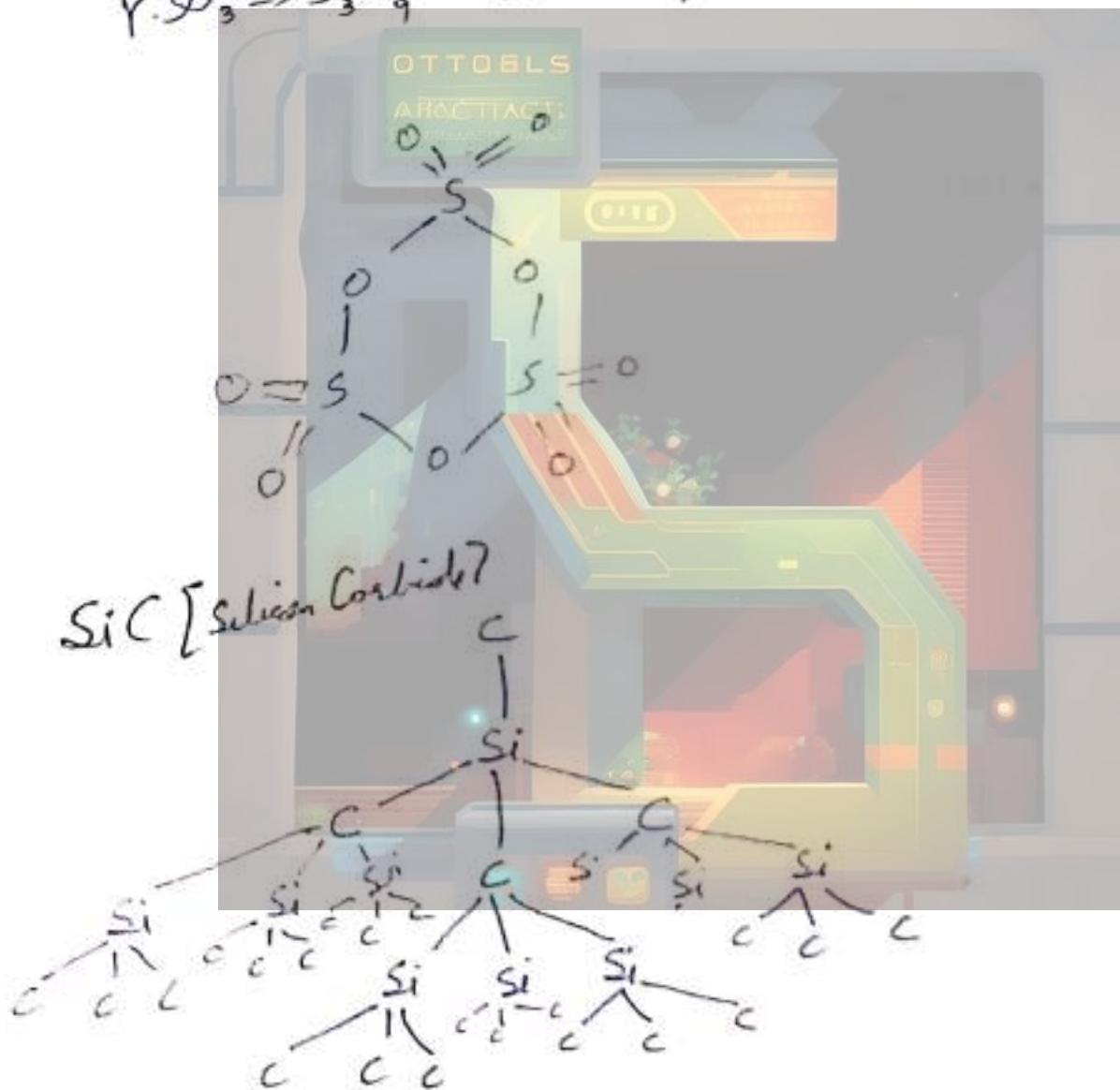
Allotropes of Tin



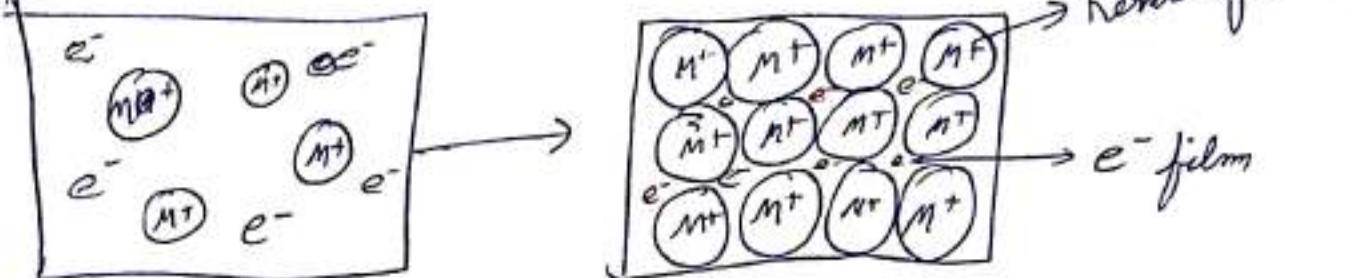
Allotropic forms of SO_3



$\gamma \cdot \text{SO}_3 \rightarrow \text{S}_3 \text{O}_9 \rightarrow$ Trimer of S^{2-}



Metallic Bond



Silicates

Beryl & Emerald

- ① Having same molecular formula but Beryl is colourless or light yellow color while emerald is green in color.



- ② In Emerald, (Cr^{+3}) is present as impurity so green in color.

Zoisite:- 3D silicate

