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11th Class Chemistry

Chemical Bonding and Molecular Structure

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CHAPTER-4

Difference B/w Electrovalent bond (Ionic) and Covalent bond.

Ionic Bond

→ Ionic bond is a bond is formed by the complete transference of a certain number of e^-'s from one atom to another atom.

e.g:- NaCl.

$Na^{\bullet} + :\ddot{Cl}: \rightarrow [Na]^{+} [:\ddot{Cl}:]^{-}$

$Na^{+} \equiv :\ddot{Cl}:$

↓

Electrostatic force of attraction.

→ They are Normally Solids.

→ There is strong force of attraction b/w ions

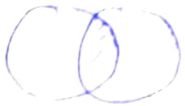
→ There Melting and boiling Points are high

→ They Conduct electricity in aqueous solution. or in the molten State

Covalent bond

→ It is formed of same or different elements combine by mutual sharing of electrons.

e.g:- H_2, O_2, HCl.



H H.

H-H.

→ They are Solid, liquid, or gas.

→ There is relatively weak force of attraction.

→ There Melting and boiling Points are very low.

→ They do not Conduct or Generally they are bad Conductors

Fajan rules [Ionic Compound]

=> In an Ionic Compound if the Polarising Power is high as well as Polarizability is high. there will be more of Covalent character less of Ionic character.

For Example :- NaCl. :- it is a ionic compound.

$Na^* \xrightarrow{+} Cl \longrightarrow Na^+ . Cl^-$

In Na^+ ($P^+ = 11$, $e^- = 10$).

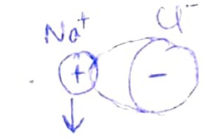
Now $P^+ > e^-$ ∴ the nuclear Power will be more.

Small Size of Cation. (Polarizing Power increases) it's nuclear Power is high

In Cl^- ($P^+ = 17$, $e^- = 18$).

Now $P^+ < e^-$ ∴ the [redacted] size will increase.

Large Size of Anion. (Polarizibility increases) it's nuclear Power is low.



Nuclear Power is higher than Cl^- nuclear Power. So it will start attract e^- from Cl. and it forms a bond like a Covalent bond.

Polarizing Power :- Tendency to attract electron cloud. [e^- cloud].

Polarisability :- The tendency of an ion to get distorted. (or Tendency of e^- cloud to bend)

Cation :- Smaller Size of cation greater Polarizing Power.

Anion :- Greater Size of Anion greater Polarizibility.

For Example :- (i) NaCl, LiCl. Here Li^+ has more Polarizing Power than Na^+.

(Size of Li^+ < Size of Na^+) ∴ LiCl has more Covalent character than NaCl.

(ii) NaCl, NaBr.

(Size of Br^- is more than Cl^-) ∴ Br^- has more Polarizability than Cl^-.

∴ NaBr has more Covalent character than NaCl.

(iii) $FeCl_2$, $FeCl_3$

Fe^{2+}, Fe^{3+}

(Size of Fe^{3+} < Fe^{2+}) ∴ Fe^{3+} has more Polarizing Power than Fe^{2+}.

∴ $FeCl_3$ has more Covalent character than $FeCl_2$.

Lattice Energy :- The energy released when one mole of ionic compound is formed from its ions in their gaseous state.

$$Na^+_{(g)} + Cl^-_{(g)} \longrightarrow NaCl_{(s)} \qquad \Delta H = L.E = -ve.$$

(Exothermic Reax).

or.

The Energy required to completely seperate a mole of solid ionic compound into gaseous ions.

$$NaCl_{(s)} \longrightarrow Na^+_{(g)} + Cl^- \qquad \Delta H = L.E = +ve.$$

(Endothermic Reax).

$$L.E = \frac{K\, q_+ q_-}{r_c + r_a}$$

where q_+ = charge of cation, q_- = " " anion, r_c = Radius of cation, r_a = " " Anion.

$L.E \propto q_+$

$Mg^{2+}Cl_2 < Al^{3+}Cl_3$

$L.E \propto q_-$

$Al_2C_3 < AlN$

C^{2-} $\quad$ N^{3-}

$L.E \propto \frac{1}{r_c}$

Size of atom (cation) decreases.

L.E increases

$LiF > NaF$.

$L.E \propto \frac{1}{r_a}$

Anion Size decrease.

L.E increases.

$KBr > KI$

Small than I.

→ Lattice energy is directly proportional to charge of ion (cation & Anion).

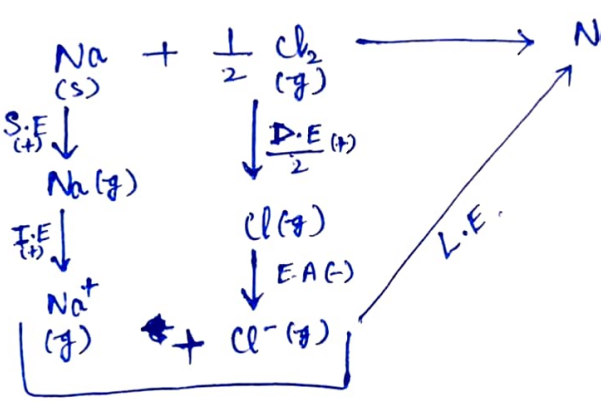
$(L.E \propto q_+ q_-)$.

→ Lattice energy is inversely proportional to Size of ions.

$$L.E \propto \frac{1}{r_c + r_a}.$$

Born Haber Cycle :- For Formation of ionic compound e.g. Na^+Cl^-.

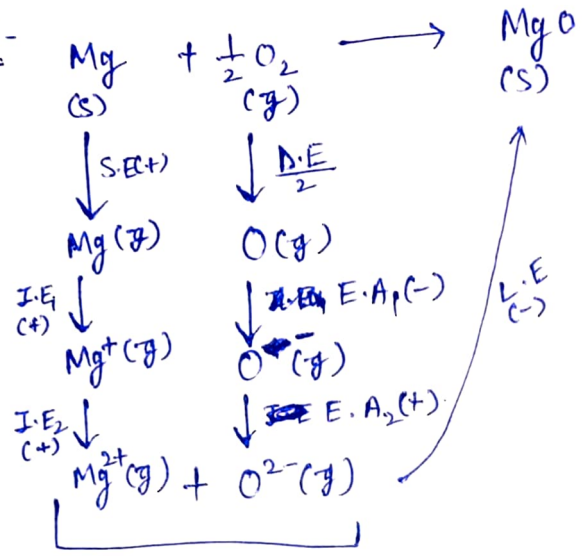
(ΔH_f = Heat of formation. (Heat Released))



$$\Delta H_f = S.E + I.E + \frac{D.E}{2} - E.A - L.E.$$

Where S.E = Sublimation energy.
I.E = Ionization energy.
D.E = Bond Dissociation energy.
E.A = Electron affinity.
L.E = Lattice energy.

Example :-

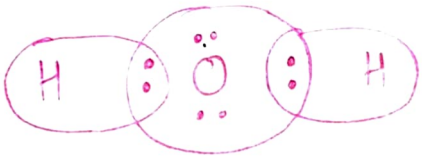


$\Delta H_f = S.E + I.E_1 + I.E_2 + \frac{D.E}{2} - E.A_1 + E.A_2 - L.E.$

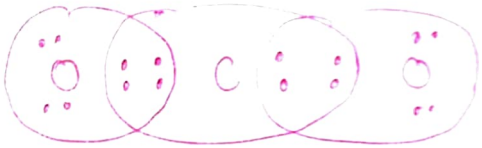
Kossel - Lewis Approach to chemical bonding :-

=> The atoms of different elements combine with each other in order to complete their octets or duplets to attain stable configuration.

Example :- (a) H_2O.



(b) CO_2.



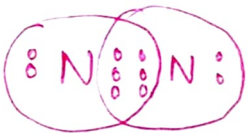
(c) N_2

=> Total No. of Valence $e^- = 5+5 = 10$

=> No. of e⁻s required to complete octet or duplet $= 2 \times 8 = 16$.

=> No. of shared $e^- = 16 - 10 = 6$.

=> No. of unshared $e^- = 10 - 6 = 4$.



or $:N \equiv N:$

Formal Charge :-

=> The formal charges is the electrical charge difference b/w the Valence e^- in an isolated atom and the Number of electrons assigned to that atom in a Lewis Structure.

=> F.C = [Total number of V.E] − [Total no. of unshared e⁻s (L.P)] − $\frac{1}{2}$ [Total no. of shared electrons (B.P)]

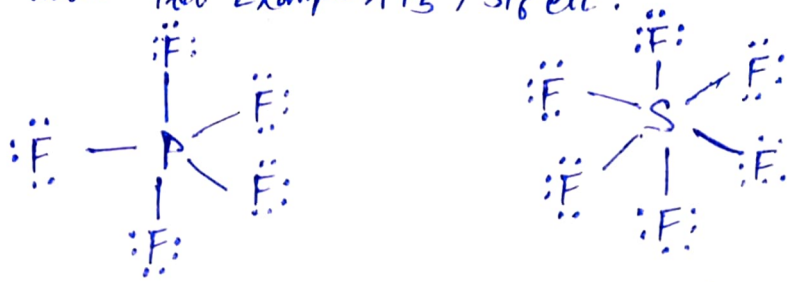
[Formal charge on an atom in a Lewis Structure].

Limitations of octet Rule :-

The Rule failed to predict the Shape and Relative Stability of molecules.

→ It is based upon the inert nature of noble gases. However, Some noble gases like Xenon and Krypton form compounds Such as XeF_2, KrF_2 etc.

→ The element present in Third period have more than eight valence electrons around the Central atom. For Example: PF_5, SF_6 etc.



→ The octet rule is not satisfied for all atoms in a molecule having an odd no. of e⁻'s for Ex. :- NO and NO_2.

→ This rule cannot be applied to those Compounds in which the no. of e⁻s surrounding the Central atom is less than eight. Eg:- LiCl, BeH_2, $AlCl_3$ etc.

Dipole Moment :-

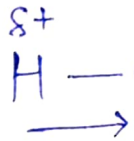
It is the Product of the magnitude of charge on anyone of the atoms of distance b/w them.

$\mu = e \times d$.

Where μ = Dipole moment.
e = charge on anyone of the atoms.
d = distance b/w the atoms.

Its unit in C.G.S System is debye (D).

Example:-



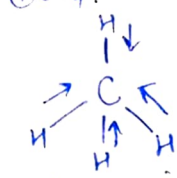
(Dipole moment has a magnitude and a Direction).

(1) => It is vector quantity and Points towards more electronegative atom.

$H \rightarrow Cl$.

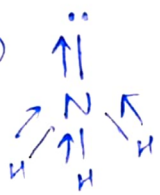
Example:-

(1) CH_4.



(Vector Sum in the same direction Hence dipole moment will be more.)

(2)



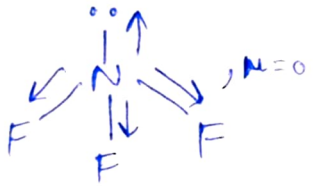
(Same Reason)

(3) $O \overleftarrow{=} C \overrightarrow{=} O$, $\mu = 0$

(4) $S \overleftarrow{=} C \overrightarrow{=} S$, $\mu = 0$.

(5) $F \overleftarrow{—} Be \overrightarrow{—} F$, $\mu = 0$

(6)



$\mu = 0$

(Vector Sum in the opposite direction, Hence dipole moment will be zero).

Bond Parameters

=> **Bond length :-** It is the average distance b/w the centres of the nuclei of two bonded atoms in a molecule.

=> **Bond length depends upon the following factors :-**

1) → Bond length decreases with increase in bond multiplicity.

$C \equiv C < C = C < C - C$.

—— Bond length increases ⟶

2) → **Size of the Atom :-** The bond length increases with increase in the size of the atom.

$H_2 > Cl_2$

H-H Cl-Cl. (Here size of H is greater than Cl).

=> **Bond angle :-** It is the average angle b/w the orbitals of the central atom containing the bonding e⁻ Pairs in the molecule. It is expressed in degree / minute / second.

For Example :- Thus H-O-H (Bond angle in H_2O is 104.5°)

=> **Bond Dissociation energy :-** The amount of energy is required to break one mole of bonds of particular type b/w the atoms of a molecule in the gaseous state.

→ **Depends upon following factors :-**

1) **Bond Multiplicity :-** Bond dissociation energy increases with increase in Bond Multiplicity.

$C \equiv C > C = C > C - C$.

⟵ Bond Dissociation energy ——

2)

or It is no of Bonds B/w two atoms in a molecules.

Bond order :- It is a measurement of the number of e⁻s involved in bonds b/w two atoms in a molecule.

→ Most of the time, bond order is equal to the no. of bonds b/w two atoms. (Exception occur when the molecule contains antibonding orbitals).

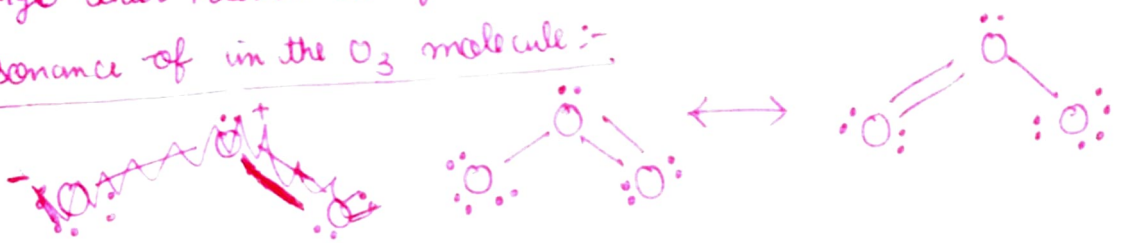
=> **Bond order** = $\left[\text{No. of Bonding } e^- s - \text{no. of antibonding } e^- s\right]$ / 2

If bond order = 0, the two atoms are not bonded.

For Example H — H , O = O , N ≡ N , C ≡ N
B.O = 1 B.O = 2 B.O = 3 B.O = 3.

Resonance :- (The Displacement of e⁻ in this way that bond change their Position to get more Stable Structure).

Resonance of in the O_3 molecule :-

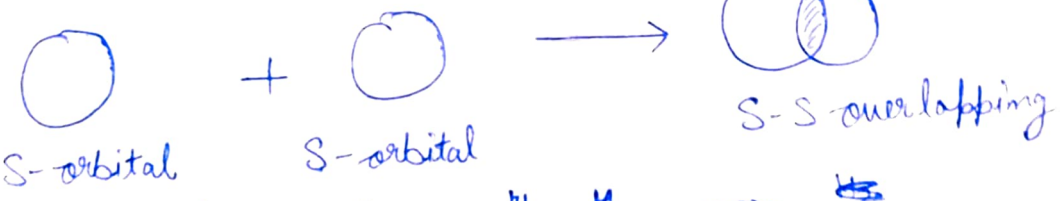


VBT (Valence Bond Theory) :- It is the overlap criteria of atomic orbitals. or [Covalent Bond is formed by overlapping of atomic orbitals].

=> **Types of overlap :-**

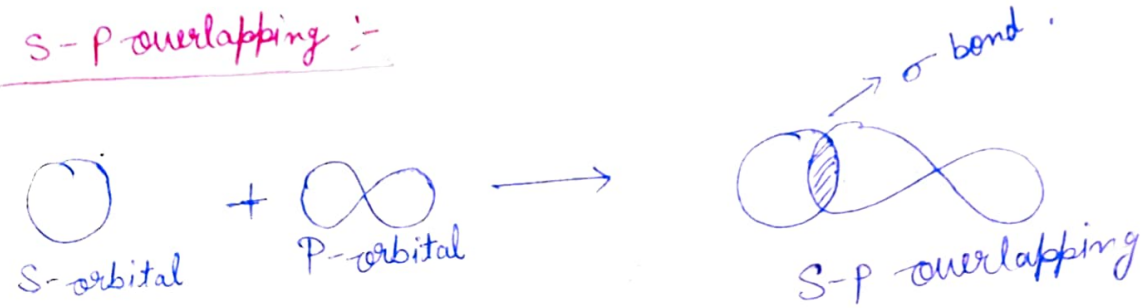
(i) **Sigma bond :-** A Covalent bond resulting from the formation of a molecular orbital by the end to end overlap of s along internuclear axis.

S - S overlapping :-



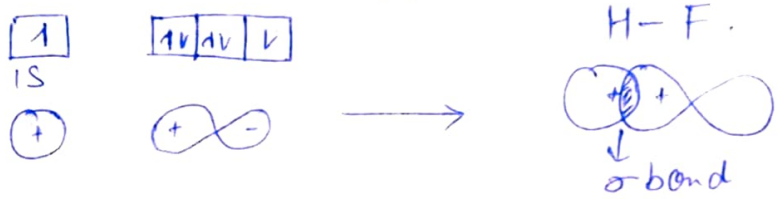
Example :- H_2 = (H — H)

(b) S-P overlapping :-



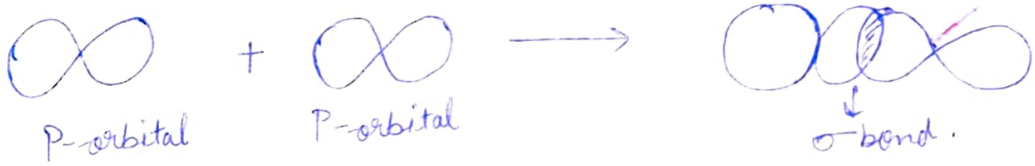
Example :-

H $\quad$ F $= 1s^2 2s^2 2p^5$

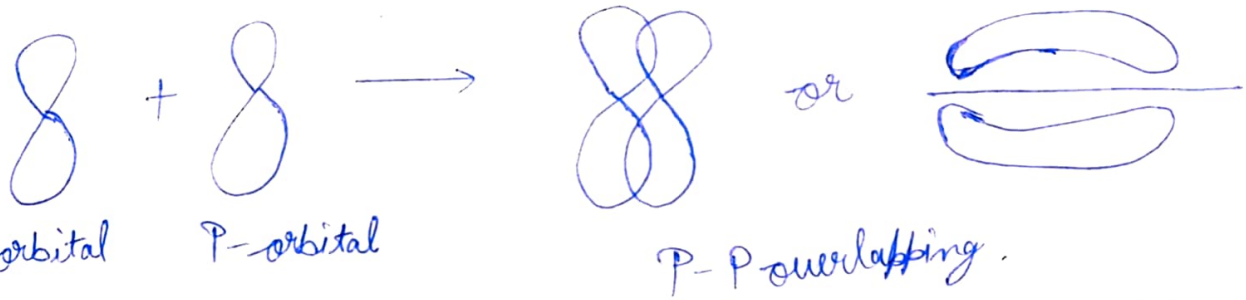


(Mostly Same Sign orbitals overlapping is Positive (Most overlap appropriate).

(C) P-P overlapping :-



(ii) **Pi (π) bond :-** A Covalent bond formation of molecular orbitals by Side-to Side overlap of atomic orbitals along a Plane Perpendicular ito a line Connecting the nuclei of the atoms.

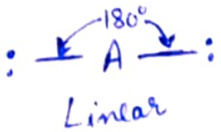
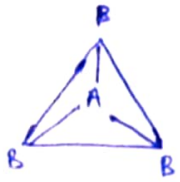
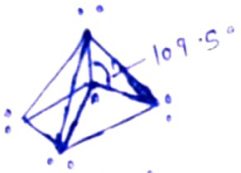
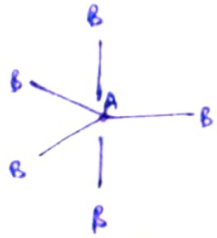
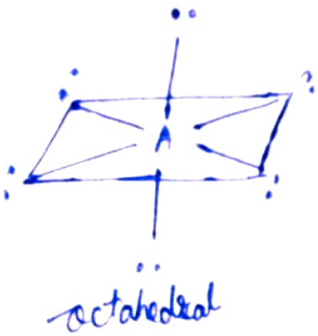
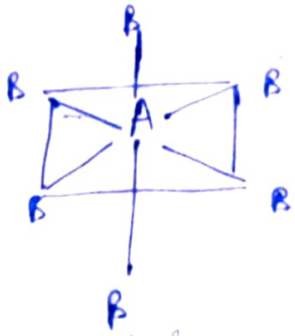


In Single bond C – C, In Double bond C $\overset{\sigma}{=}$ C (σ, π), In Triple bond C $\overset{\sigma}{\underset{\pi}{\equiv}}$ C (σ, π, π)

[Note :- σ bond is Stronger bond than π-bond (Because it overlapps internuclear axis).]

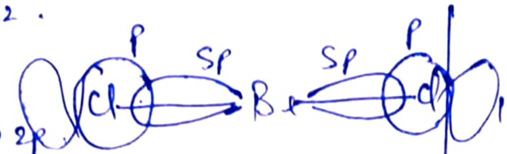
Hybridisation :- Inter mixing of Atomic orbitals of the same atom with nearly same energy, resulting in formation of new orbitals of equal and identical Shape.

Table :- Geometry of Molecules in which the Central atom has no Loan Pair of electrons.

Number of e⁻ Pairs	Arrangement of e⁻ Pairs	Molecular geometry	Examples	Hybridisation
2	 Linear	B — A — B Linear	$BeCl_2$, $HgCl_2$.	SP
3	 Triangonal Planar	 Trigonal Planar	BF_3	SP^2
4	 Tetrahedral	 Tetrahedral	CH_4, NH_4^+	SP^3
5	 Trigonal bipyramidal.	 Trigonal bipyramidal	PCl_5.	SP^3d
6.	 octahedral	 octahedral	SF_6.	SP^3d^2.

⇒ Calculate Z = No. of σ bond + L.P of Central Atom.

1) $BeCl_2$.



Be.
$1S^2 2S^2 2p$

[↿] [↿| |]

Sp.

[↿|↿]

Sp Sp.

Cl σ B σ Cl

Cl.

[⥮|⥮|↿]

Cl

[⥮|⥮|↿]

Cl —σ— Be —σ— Cl.

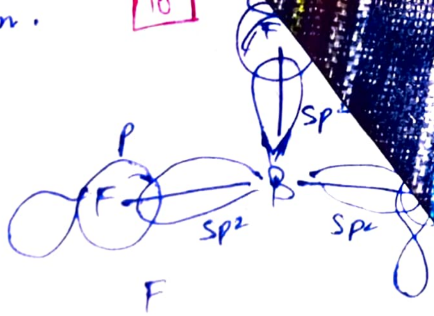
No. of σ bond is 2.

Z = 2.

→ Then Hybridisation is Sp.

→ Geometry is Linear.

2) BF_3.



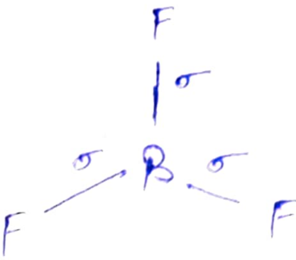
B
$1S^2 2S^2 2p^1$

[↿] [↿|↿|]

Sp^2

[↿|↿|↿]

Sp^2 Sp^2 Sp^2



No. of σ bond is 3.

Z = 3.

→ Then Hybridisation is Sp^2.

→ Geometry is Trigonal Planar.

3) CH_4.

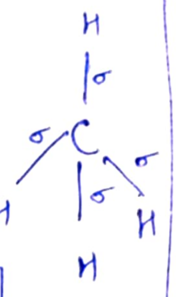
C
$1S^2 2S^2 2p^2$

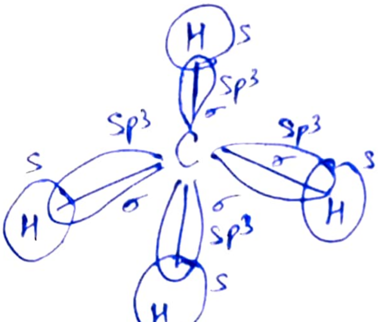
[↿] [↿|↿|↿]

Sp^3

[↿|↿|↿|↿]

Sp^3 Sp^3 Sp^3 Sp^3
S S S S





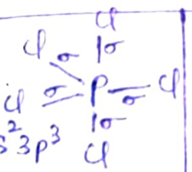
No. of σ bond is 4.

Z = 4.

Then Hybridisation is Sp^3.

Geometry is Tetrahydral.

4) PCl_5



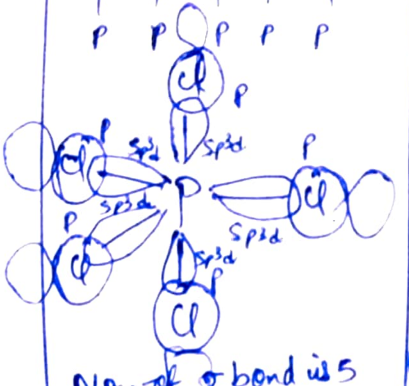
P
$1S^2 2S^2 2p^6 3S^2 3p^3$

[↿] [↿|↿|↿] [↿| | | |]

Sp^3d

[↿|↿|↿|↿|↿|]

Sp^3d Sp^3d Sp^3d Sp^3d Sp^3d
P P P P P



No. of σ bond is 5

Z = 5.

Then Hybridisation is Sp^3d.

Geometry is Triagonal Bipyramidal.

5) SF_6.

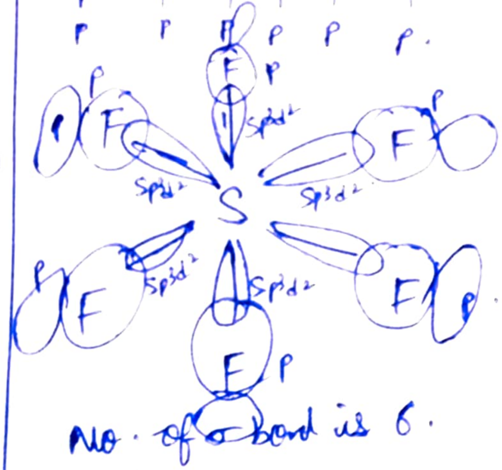
S
$1S^2 2S^2 2p^6 3S^2 3p^4$

[↿] [↿|↿|↿] [↿|↿| | | |]

Sp^3d^2

[↿|↿|↿|↿|↿|↿]

Sp^3d^2 Sp^3d^2 Sp^3d^2 Sp^3d^2 Sp^3d^2 Sp^3d^2.
P P P P P P.



No. of σ bond is 6.

Z = 6.

The Hybridisation is Sp^3d^2.

Geometry is octahedral.

Hybridisation of $BeCl_2$.

⇒ Be (Z = 4)

$1s^2 2s^2 2p$.

(Ground state) [↑↓] 2s [| |] 2p

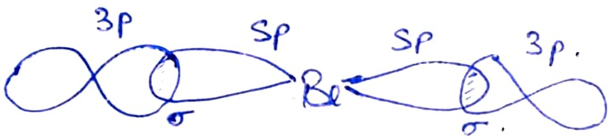
(Excited state) [↑] [↑| |]

New Hybrid orbitals [↑|↑]
sp sp
(3p) Cl Cl (3p).

Cl.
$1s^2 2s^2 2p^6 3s^2 3p^5$
[↑↓|↑↓|↓]
3p.

Cl —σ— Be —σ— Cl

There are σ bonds = 2.
Hybridisation is sp.
Geometry is Linear.



Hybridisation of PCl_5.

⇒ P (Z = 15)

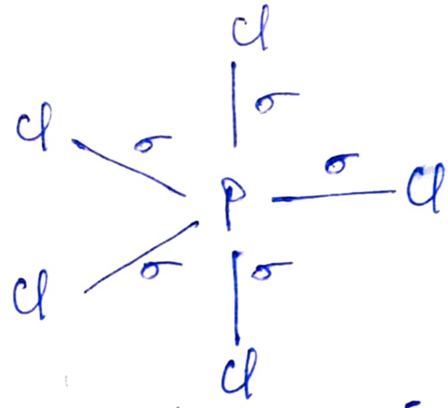
$1s^2 2s^2 2p^6 3s^2 3p^3$

(Ground State of P) [↑↓] 3s [↑|↑|↑] 3p

(Excited state of P) [↑] 3s [↑|↑|↑] 3p [↑| | | |] 3d

New Hybrid orbitals. [↑|↑|↑|↑|↑]
sp^3d sp^3d sp^3d sp^3d sp^3d
(3p) Cl (3p) Cl (3p) Cl (3p) Cl (3p) Cl.

Cl.
$1s^2 2s^2 2p^6 3s^2 3p^5$.
[↑↓|↑↓|↓]



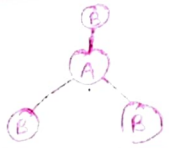
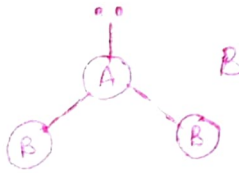
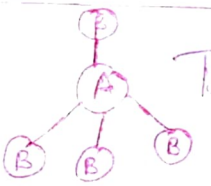
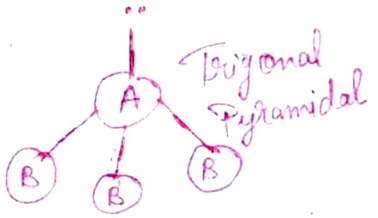
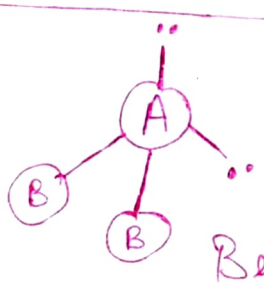
No. of σ bonds are 5.
Then Hybridisation is sp^3d.
Geometry is Trigonal bipyramidal

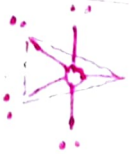
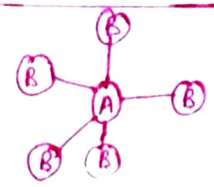
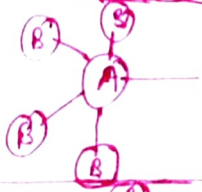
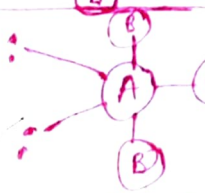
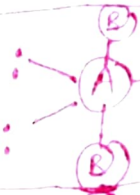
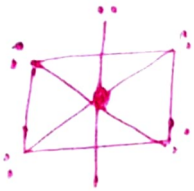
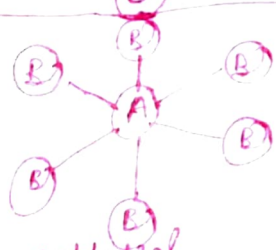
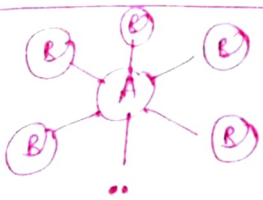
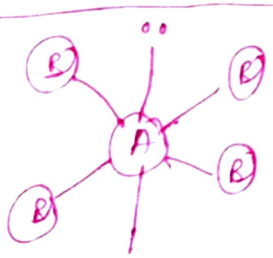
VSEPR (Valence Shell electron Pair repulsion theory).

=> There are three types of repulsion in a covalent molecule
lp:-lp > lp:-b.p > b.p - b.p.

=> These repulsion depend upon difference in electronegativity between A and B.

=> The hybridisation of central atom decides geometry and VSEPR theory the shape of molecule.

Total No. of e⁻ Pairs	Geometry of the e⁻ Pairs.	B.P | L.P (Molecule formula)	Geometry (Shape) of the molecule	Examples
2 Hybridisation = Sp	:—○—: Linear.	2 | 0 (AB_2)	 Linear	BeF_2, $BaCl_2$, $HgCl_2$.
3 H = Sp^2	 Triangular Planar	3 | 0 (AB_3)	 Triangular Planar	BF_3; $AlCl_3$, SO_3.
		2 | 1 (AB_2L)	 Bent (V-Shape)	SO_2, O_3, $SnCl_2$, NO_2
4 H = Sp^3	 Tetrahedral	4 | 0 (AB_4)	 Tetrahedral	SiF_4, NH_4^+ CH_4.
		3 | 1 AB_3L	 Trigonal Pyramidal	NH_3, PCl_3, NCl_3, NF_3, PH_3, H_3O^+
		2 | 2 AB_2L_2	 Bent	H_2S; NH_2^-, H_2O.

Hybridisation of the C-Pairs.	B.P	L.P.	(Molecule formula)	Geometry (Shape) of the molecule.	Example.
sp^3d  Trigonal bipyramidal	5	0	(AB_5)	 Trigonal bipyramidal	PCl_5. PF_5, & $SbCl_5$,
	4	1	(AB_4L)	 See Saw	SF_4.
	3	2	(AB_3L_2)	 T-Shaped	ClF_3.
	2	3	(AB_2L_3)	 Linear	XeF_2.
6 $H = sp^3d^2$  octahedral	6	0	(AB_6)	 octahedral	SF_6.
	5	1	(AB_5L)	 Square pyramidal	ClF_5.
	4	1	(AB_4L_2)	 Square planar	XeF_4.

Example :-

(i) SO_2.

S (Z=16). $1s^2 2s^2 2p^6 3s^2 3p^4$

O (Z=8). $1s^2 2s^2 2p^4$

Number of e⁻ Pair = Number of σ Bond + Lone Pair

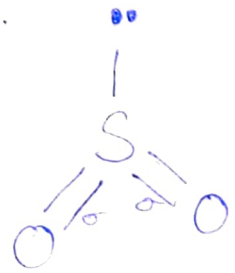
= 2 + 1 = 3.

Hybridisation = Sp^2

Molecular formula = AB_2L

Geometry = Triangular Planar.

Shape = Bent (V-Shape).



Rough

S (V.E) = 6.

O O O O O O

O O 1L.P.

AB_2L

(ii) PCl_3.

P (Z=15) $1s^2 2s^2 2p^6 3s^2 3p^3$

Cl (Z=17). $1s^2 2s^2 2p^6 3s^2 3p^5$

No. of e⁻ Pair = No. of σ Bond + Lone Pair

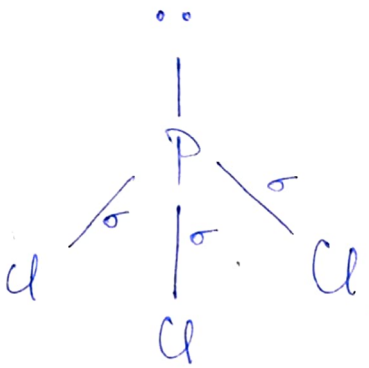
= 3 + 1 = 4.

Hybridisation = Sp^3

Molecular formula = AB_3L

Geometry = Tetrahedral.

Shape = Trigonal Pyramidal



O O O O O

Cl Cl Cl 1L.P.

AB_3L.

Hybridisation in C_2H_6 molecule. (Ethane)

C (Z = 6).

$1s^2 2s^2 2p^2$

Ground State: 1↓ | 1 | 1 | | (2s | 2p)

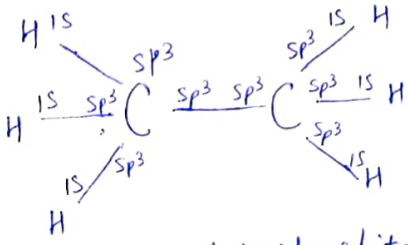
Excited State: 1 | 1 | 1 | 1 — sp^3

New hybrid orbitals: sp^3 sp^3 sp^3 sp^3

H (Z = 1).

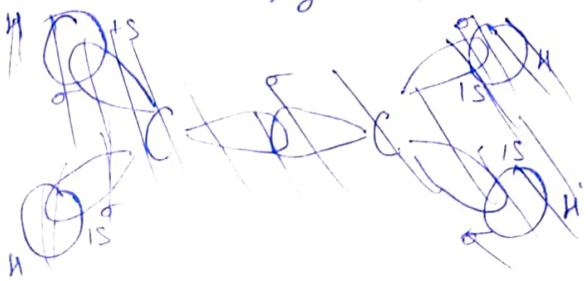
$1s^1$

1



unhybrid orbital = 0.

Hybrid orbital = 4.



Hybridisation in C_2H_4 molecule. (Ethene)

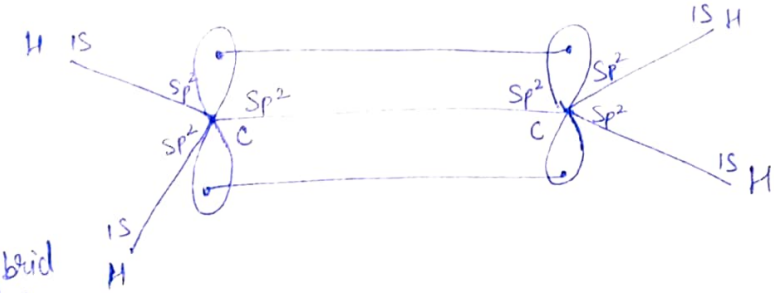
$CH_2 = CH_2$.

C (Z = 6).

$1s^2 2s^2 2p^2$.

Ground State: 1↓ | 1 | 1 | |

Excited State: 1 | 1 | 1 | 1 — sp^2 | unhybrid orbital.



Hybridisation of C_2H_2 molecule. (Ethyne).

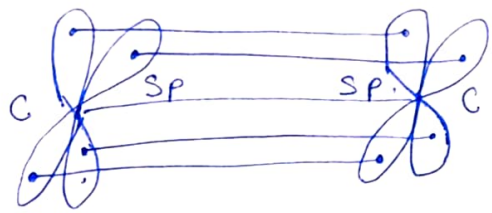
$HC \equiv CH$.

C (Z = 6).

$1s^2 2s^2 2p^2$

Ground State: 1↓ | 1 | 1 | |

Excited State: 1 | 1 | 1 | 1 — Sp | unhybrid orbitals.



Atomic orbital (V.B.T)

Atomic orbital is the region having the highest probability of finding an e^- in an atom.

→ Schrodinger equation is used.

→ It fails to explain Paramagnetic character of O_2.

→ It fails to explain the existance of H_2^+

Molecular orbital

→ Molecular orbital is the region having the highest probability of finding an electron of a molecule.

→ Linear Combination of atomic orbitals (LCAO) is used.

→ It explains the Para magnetic character of O_2

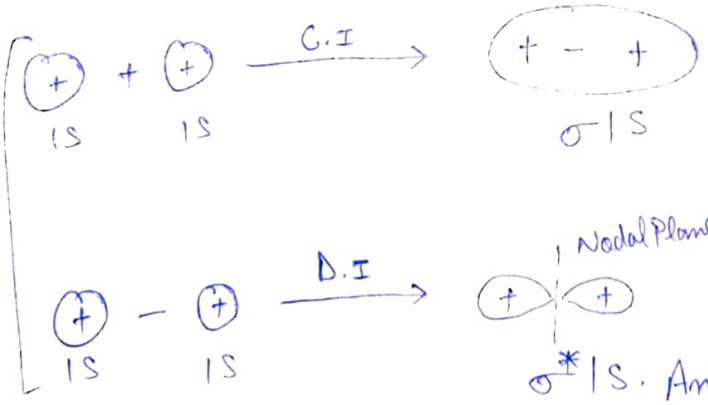
→ It explains the existance of H_2^+.

Formation of Bonding molecular orbital and Antibonding molecular orbital.

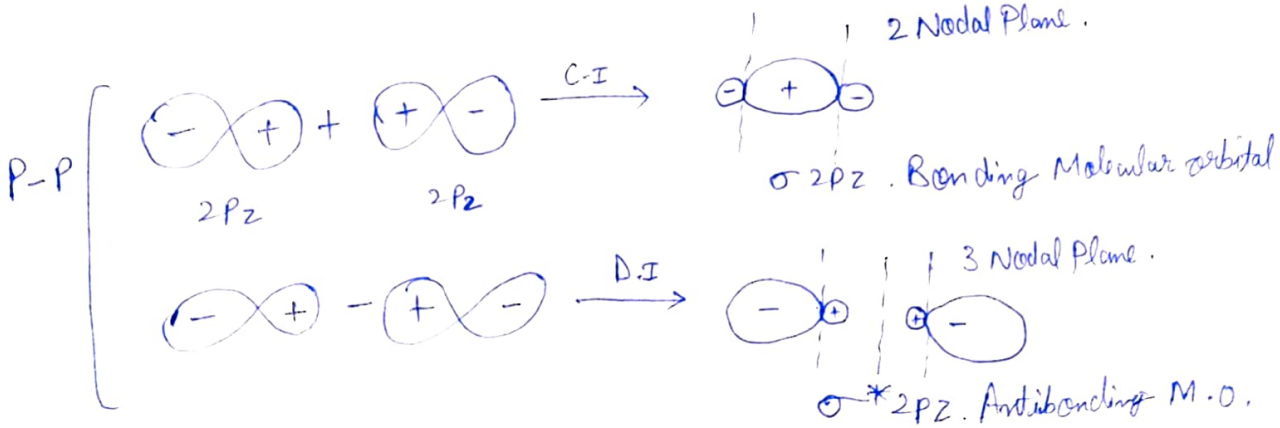
[we have two (σ, π) Molecular orbitals].

⇒ σ Molecular orbital :- Axial Atomic orbital overlapping.

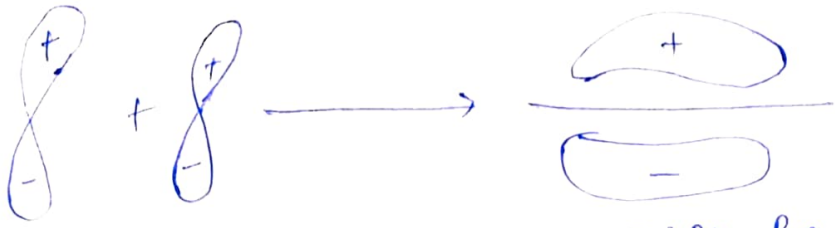
S-S



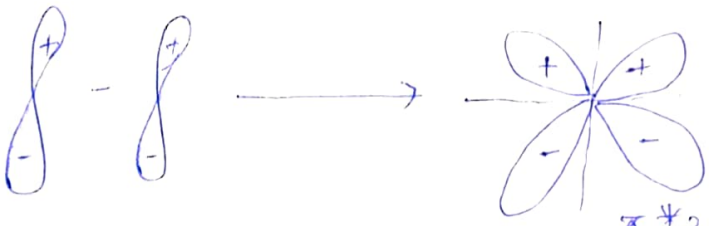
P-P



π Molecular orbital :- (When Sidewise overlapping of A



$\pi 2p_x$ Bonding M.O.



$\pi^* 2p_x$ Antibonding M.O.

Bonding Molecular orbital

→ These orbitals have less energy.

→ They are More Stable.

→ Like :- $\sigma 1s$, $\sigma 2p_z$, $\pi 2p_x$, $\pi 2p_y$ etc

Anti Bonding Molecular orbital

→ These orbitals have high energy.

→ They are less Stable.

→ orbitals like :- $\sigma^* 1s$, $\sigma^* 2p_z$ etc.

How to do e⁻ Configuration of Molecules :-

→ Same Rules follow to fill the electrons in orbitals.

(a) Afbau Principle. (b) Pauli's Exclusion, (c) Hund's Rule.

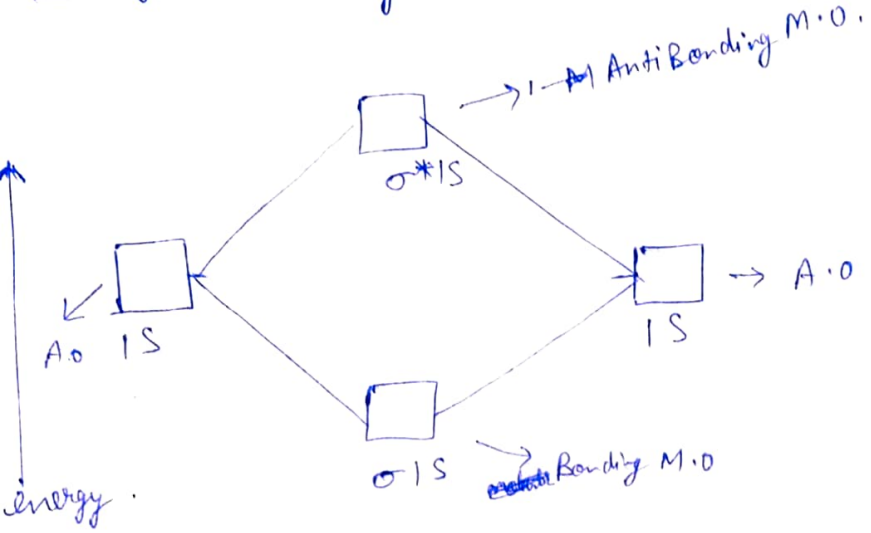
↙ Here lower energy orbital is Bonding Molecular orbital

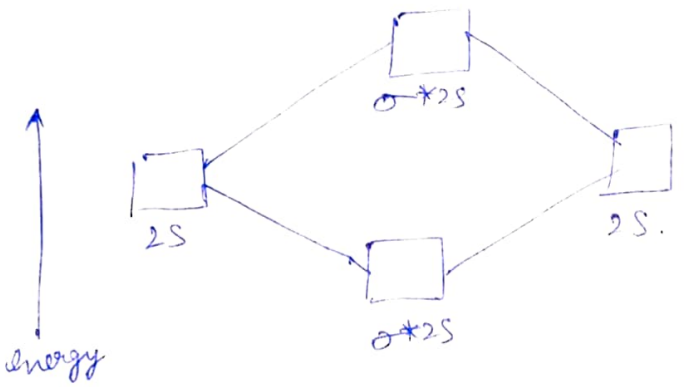
↘ and Higher energy orbital is Anti Bonding Molecular orbital.

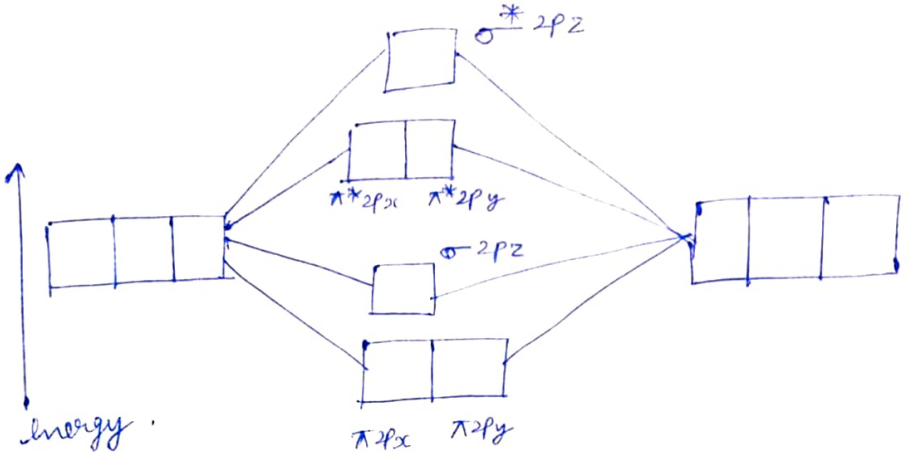
→ The electronic Configuration is for the molecule which have (less than or equal to 14 e⁻ or ~~=~~ ≥ 14).

$(\sigma 1s\ \sigma^* 1s)\ (\sigma 2s\ \sigma^* 2s)\ (\pi 2p_x = \pi 2p_y)\ \sigma 2p_z\ (\pi^* 2p_x = \pi^* 2p_y)\ \sigma^* 2p_z$.

Energy level diagram :- $\geq 14e^-$







For Example :-

E.C of Li_2. (No. of e^- = 6). → Paired e^-.

$(\sigma 1s^2)(\sigma^* 1s^2)(\sigma 2s^2)$ ∴ it is diamagnetic character.

Li_2^+ (No. of e^- = 7). → B.O = $\frac{N_b - N_a}{2} = \frac{4-2}{2} = 1$

$(\sigma 1s^2)(\sigma^* 1s^2)(\sigma 2s^2)(\sigma^* 2s^1)$ → unpaired e^-.

(Magnetic Nature ⇒ it is Paramagnetic).

→ B.O = $\frac{N_b - N_a}{2} = \frac{4-3}{2} = \frac{1}{2}$.

Li_2^+ (No. of e^- = 5).

$(\sigma 1s^2)(\sigma^* 1s^2)(\sigma 2s^1)$ → unpaired e^-.

∴ it is Paramagnetic.

→ B.O = $\frac{N_b - N_a}{2} = \frac{3-2}{2} = \frac{1}{2}$

B.O ∝ Stability.

[Note :- If we get Same B.O of the molecule. then we have to check it's antibonding e⁻'s. if antibonding e⁻ will be less. it will be more Stable].

$Li_2 > Li_2^+ > Li_2^-$

← Stability increases.

He molecule (He_2).

(No. of e^- = 4).

$(\sigma 1s^2)(\sigma^* 1s^2)$. B.O = $\frac{N_b - N_a}{2} = \frac{2-2}{2} = \frac{0}{2} = 0$.

He molecule does not exist.

(if any molecule B.O will be zero the molecule will not exist).

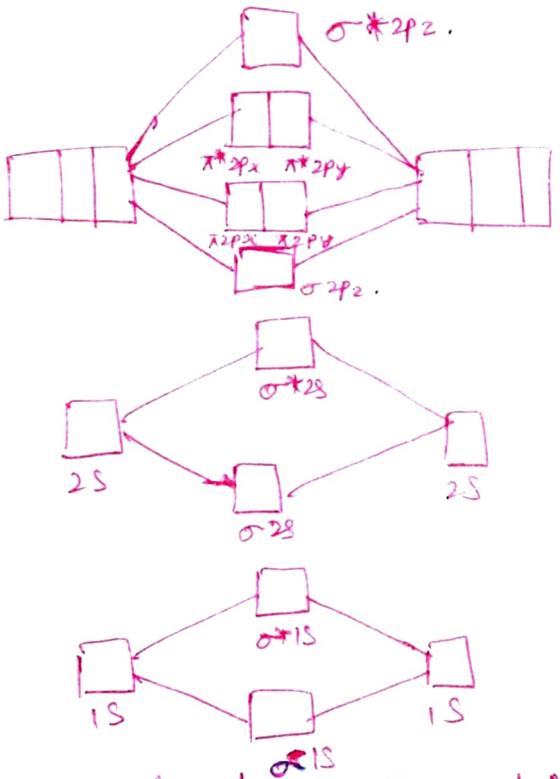
If molecule contain $e^- > 15$ and < 20.
$(15-20)$..

Then, the E.C will be.

$(\sigma 1s\ \sigma^* 1s)(\sigma 2s\ \sigma^* 2s)\ \sigma 2pz\ (\pi 2px = \pi 2py)\ (\pi^* 2px = \pi^* 2py)$

Energy level diagram

$\sigma^* 2pz$.



Write e⁻ Configuration :- O_2 and N_2.

⇒ i) O_2. (Total No. of $e^- = 16$).

$(\sigma 1s^2\ \sigma^* 1s^2)\ (\sigma 2s^2\ \sigma^* 2s^2).\ \sigma 2pz^2\ (\pi 2px^2 = \pi 2py^2)\ (\pi^{*1}_{2px} = \pi^{*1}_{2py})$

$(KK)\quad (KK)'$

→ Unpaired electrons.
→ Magnetic Nature :- Paramagnetic
→ $B.O = \frac{N_b - N_a}{2} = \frac{10-6}{2} = \frac{4}{2} = 2$.

⇒ N_2 (Total No. of $e = 14$).

$(\sigma 1s^2 \sigma^* 1s^2)\ (\sigma 2s^2\ \sigma^* 2s^2)\ (\pi 2px^2 = \pi 2py^2)\ \sigma 2pz^2$

→ Paired electrons, → Magnetic Nature = Diamagnetic

→ $B.O = \frac{N_b - N_a}{2} = \frac{10-4}{2} = \frac{6}{2} = 3$.