3. CHEMICAL BOND

Electronic - Valency theory:-

In periodic table Inert gas elements are having a stable electronic configuration. Except Helium (He) the remaining elements has got ns^2np^6 configuration known to be octet configu. the remaining elements are not at all having this type of configuration. So that all the elements elements ready to achieve this configuration by loosing electrons (OL) by gaining electrons (GL) Sharing electrons is known to be electronic - valency theory.

* Ionic Bond:-

Due to the migration of electron(s) from one atom to another atom cation and Anion results to give an Ionic bond by Strong Electro-Static attraction forces

Ex:- (i) Sodium chloride NaCl:-

Atomic No. of Sodium (Na) = 11

It's electronic configuration $= 1s^2 2s^2 2p^6 3s^1$

To attain Stability 'Na' atom give up one electron by I^{st} Ionisation Enthalpy.

$$Na \xrightarrow{I.E} Na^{\oplus} + e^{\ominus}$$

Atomic number of chlorine (Cl) = 17

Valency configuration $= 3s^2 3p^5$

= ↿⇂ ↿⇂ ↿⇂ ↑

To attain stability 'Cl' atom up one electron to get stable configuration

$$Cl + e^{\ominus} \longrightarrow Cl^{\ominus} \quad (Anion)$$

Then, Due to strong electrostatic attraction forces 'Na+' and 'Cl' combines to give NaCl.

$$Na^{\oplus} + Cl^{\ominus} \longrightarrow NaCl$$

(ii) $AlCl_3$

Atomic no. of Al = 13

Valency configuration = $2s^2 2p^6 3s^2 3p^1$

To attain stability Aluminium atom give up '3' electrons.

$$Al \longrightarrow Al^{+3} + 3e^{\ominus}$$

Atomic no. of chlorine (Cl) = 17

Valency configuration = $2s^2 2p^6 3s^2 3p^5$

To attain stability 'Cl' atom up '8' electro to get stable configuration

~~Cl + e = Cl~~ (Empty)

$$Cl + e = Cl^{\ominus}$$
$$Cl + e = Cl^{\ominus}$$
$$Cl + e = Cl^{\ominus}$$
$$3Cl + 3e = 3Cl^{\ominus}$$

Aluminium combines with 3 chloride ions to give an ionic compound aluminium chloride

* Lewis dot method:

Na = 11

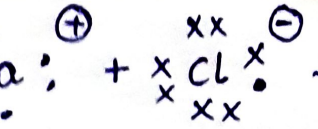
Valency configuration = $2s^2 2p^6 (3s^1)$

$$Na \longrightarrow Na^{\oplus} + e^{\ominus}$$

Lewis Dot method = $Na \longrightarrow :\dot{Na}: + \cdot(e^-)$

Chlorine = $3s^2 3p^5$ => [1L] [1L 1L 1]

$$Cl + \cdot e^{\ominus} \longrightarrow$$

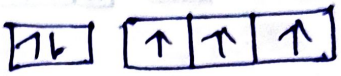


$\longrightarrow$ NaCl (Ionic bond)

* (ii) NF_3

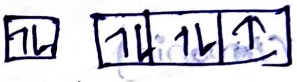
Nitrogen = 7

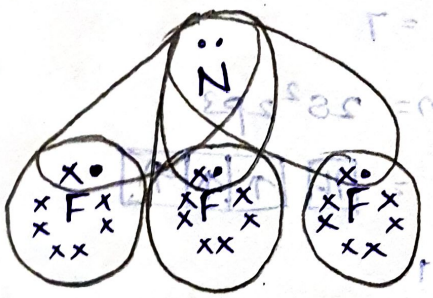
Valency configuration = $2s^2 2p^3$

= 

F = 9

Valency configuration = $2s^2 2p^5$

= 



<u>Covelent bond</u>:-

The sharing of electron pair(s) b/w two atoms is known to be covelent bond.

<u>Ex</u>:- (i) H - Cl (Hydrogen chloride)

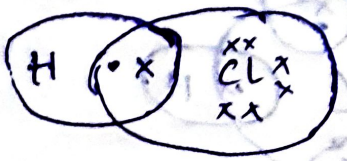
Atomic number of Hydrogen = 1

configuration = $1s^1$

Atomic no. of chlorine = 17

Configuration = $3s^2 3p^5$

= 



H - Cl

(ii) H_2O

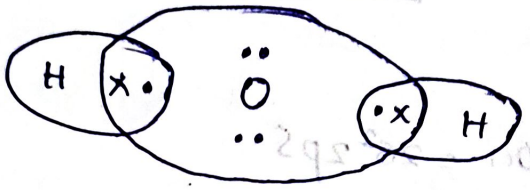
Atomic no. of Hydrogen = 1

Configuration = $1s^1$

Atomic no. of oxygen = 8

Configuration = $2s^2 2p^4$

= 



(iii) NH_3 (Ammonia)

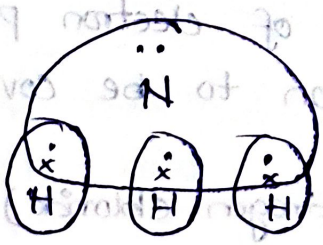
Atomic no. of Nitrogen = 7

Electronic configuration = $2s^2 2p^3$

= 

Atomic no. of (H) = 1

valency configuration = $1s^1$

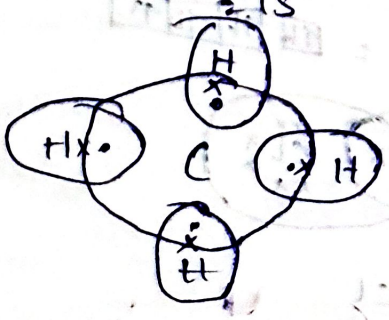


(iv) CH_4 (Methene)

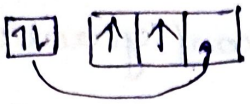
Atomic no. of Carbon = 6, configuration = $2s^2 2p^2$

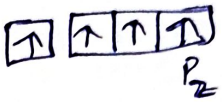
Atomic no. of Hydrogen = 1

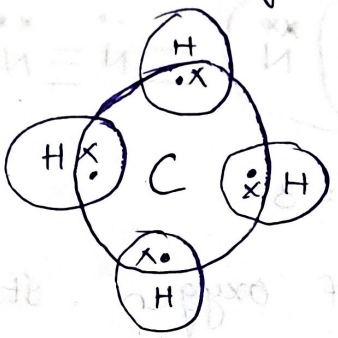
= $1s^1$



According to ground State configuration, carbon atom has two unpaired electrons only, That's why it is capable to form two bonds only. In reality carbon forms '4' bonds with 4 Hydrogen atoms (CH_4). So, ground State configuration is not applicable. In the presence of high energy the $2s^2$ electrons are exicted to vacant P_z orbital to give '4' unpaired electrons

$C = 2s^2 2p^2$  - vacant P_z

Exicted electronic configuration = 



Properties of covelent bond:-

) Bond distance:-

In the case of covelent bonded molecules the half of the distance b/w two centers of two Nuclie is known to be covelent radius

$H_2 + Cl_2 \rightarrow 2HCl$

In the case of Hydrogenchloride

H-Cl. The bond distance is given by 1.275 pm This is not at all consider as atomic radius of Individuals.

In the case of Homo-Neuclear molecules (CH_2, Cl_2, N_2, F_2) The covelent radius said to be Individial atomic radii

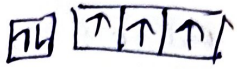
② **Bond order:-**

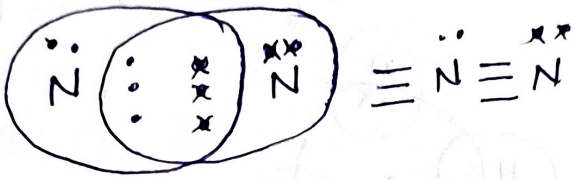
In the case of di-atomic molecules the no. of bonds placed b/w two atoms is known to be Bond order.

Ex:- i) In the case of Nitrogen (N). Its Bond order is '3'

Nitrogen atomic number = 7

Electronic configuration = $2s^2 2p^3$

= 

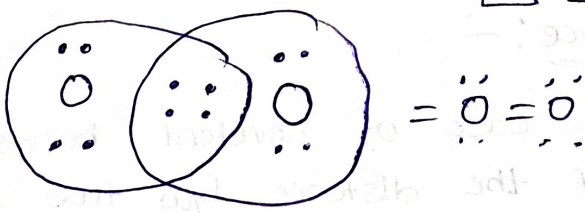


Bond order = 3

ii) In the case of oxygen. Its Bond order is 2.

atomic number of oxygen = $2s^2 2p^2$

= 



Bond order = 2

③ **Bond Angle:-**

In given molecule the practicle calculation of Bond angle towards adjecent orbitals is known to be bond angle. At this stage all the repulsions are minimised in Molecule.

Ex:- (i) In water molecule bond angle is "104°30'"

(ii) In Ammonia molecule bond angle is "107°"

Bond Dissociation Energy:-

The energy required to break one mole of bonds is known to be Bond Dissociation Energy.

Ex:- (i) Bond Dissociation of Hydrogen molecule (H_2)

$$H-H \longrightarrow \dot{H} + \dot{H} \qquad \Delta H = +435 \text{ KJ/Mole}$$

(1 Mole of Hydrogen) [Breaking]

(ii) (O_2) $O=O \longrightarrow 2\ \ddot{O}\cdot$ $\quad \Delta O = +490$ KJ/Mole

(iii) (N_2) $N\equiv N \longrightarrow 2N$ $\quad \Delta N = +900$ KJ/Mole.

~~Polar and~~

Types of covelent bond:-

Non polar covelent bond:-

If covelent bond placed b/w two similar atoms. Then, the bond pair always between two centers of two atoms. That's why the molecule doesn't attain positive & negetive charges (Di-pole)

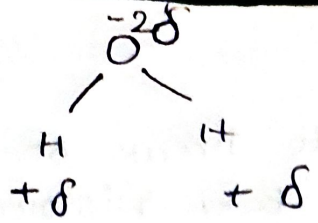
Ex:- H_2, Cl_2, N_2, O_2, F_2

Polar covelent bond:-

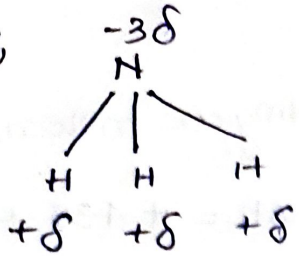
When covelent bond is placed b/w two dissimilar atoms ("Hetero"). Then, the bond pair of electrons shift towards (Partially) one particular atom which has got high electronegetively. then this atom attains partial -ve charge. and other atom gets partial +ve charge. Due, to the formation of this charges the covelent bond is known to be Polar covelent bond.

Ex:- $\overset{+\delta}{H}-\overset{-\delta}{Cl}$

(ii)



(iii)



(iv)

Di-pole moment :-(μ) :-

Di-pole moment is a product of charge and distance b/w two poles. therefore, Di-pole

Di-pole moment $\boxed{\mu = \delta . l}$

= charge . Distance

$\therefore$ μ = colomb meter units

The units for Dipole moment is 'DEBYE'

1 Debye = 1D = 3.33×10^{-30} coloumb meter.

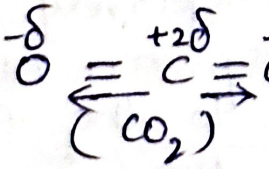
Ex :- (i) 1.875 D

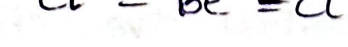
It means $1.875 \times 3.33 \times 10^{-30}$ colomb meter

(ii) 2D

→ $2 \times 3.33 \times 10^{-30}$

In the case of linear molecules like carbon dioxide or (CO_2) and $BeCl_2$

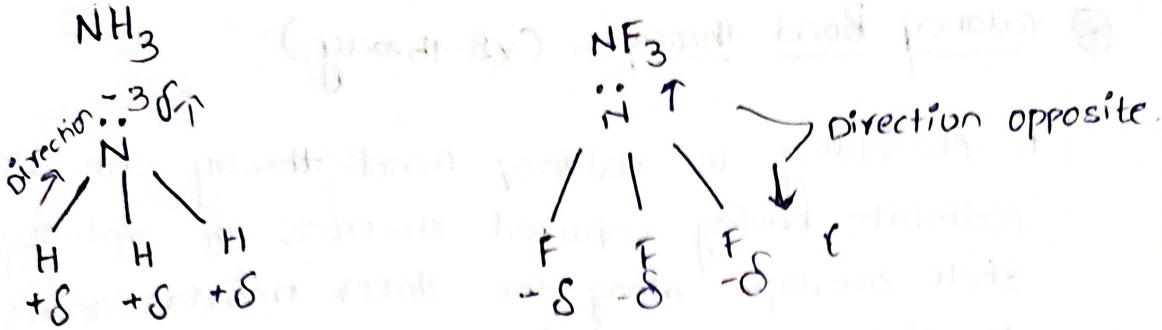




Since, Dipole moment is a vector quantity which is a product of magnitude and Direction that's why in the case of linear molecules equal & opposite Di-pole moment

values in algebric sum is said to be zero.

* Discuss the Di-pole moment of NH_3 and NF_3:-



In the case of Ammonia 'N' has got hight Electronegitivity. That's why the bond pair of electrons moves towards nitrogen. and the same time the loan pair on 'N'-atom (is) also ~~moves~~ tries to move in the same direction (forward) so that the bond pair of electrons in Ammonia and the loan pair travelles is same direction. so that the molecule is having good Di-pole moment.

~~In~~ In contrast to this, In the case of NF_3 molecules. The Bond pair of electrons moves towards Florine atom and the same time. The loan pair of electrons on Nitrogen atom moves Forward direction. this results a disturbance in charge and hence, the calculated Di-pole moment is higher for NH_3 (Ammonia) than NF_3 molecule.

(4m) ~~****~~

* Fajans Law:-

1. (Fajans Law), in given cations, which cation has got low atomic radii comes under covelent character.
2. In given cations which one is having high Ionization Enthalpy are favourable for covelent bond.
3. In given cations which is having high charge comes under covelent character,
4. In given cations which cation has got "Pseudo Inert gas configuration comes under covelent character.

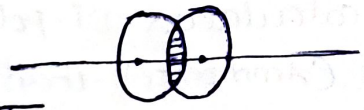
Similarly, we can explain covelent bond for anions.

⊛ Valancy Bond theory:- (VB theory)

1. According to valency Bond theory the atomic oribitals having unpaired electrons of valency shell overlaps along the Inter-necular access to give covelent bond.
2. As the over lapping increases strong covelent bond results as the overlapping is minimum weak covelent bond oresults.
3. If the over lapping of p-orbitals takes place Perpendicular to the Inter-necular access to give a weak 'π'-bond takes place. on the Other hand the strong covelent bond oresults on Internuclear access to give "sigma" (σ) bond takes places.

Ex: ① H_2 molecule-

According to VB theory the over lapping of '2' 1S orbital along the Inter-nuclear axes gives covelent bond



σ_{S-S} S-S over lapping (σ_{S-S} bond)

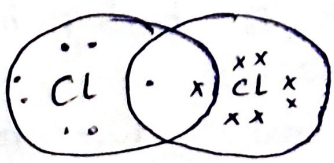
- [Internucler axis:- means an Imaginary line Passes through 2 centers of 2 nuclie]

② Cl_2 molecule:-

chlorine Atomic no. = 17

Electronic configuration = $3s^2 3p^5$

= $\boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\ \boxed{\uparrow\downarrow}\ \boxed{\uparrow}$
$P_x\ P_y\ P_z$



$\equiv$ Cl-Cl

∞ + ∞

σ_{P-P}

σ_{P-P} (P-P over lapping)

(3) HCl molecule :-

H = 1 = $1s^1$

O + ∞

σ_{S-P}

σ_{SP} (S-P over lapping)

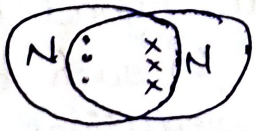
During the formation of covelent bond two atomic orbitals placed with single electrons overlapps along the Internuclear axis oresults "sigma Bond(σ)" it is a strong Bond because the newly formed electron pair placed b/w two centers of two Nuclie

4) N_2 molecule :-

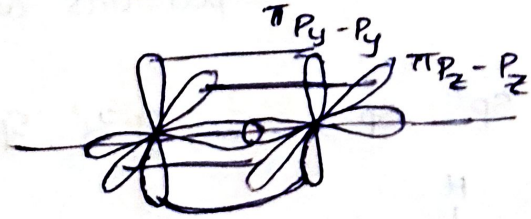
Atomic number of Nitrogen = 7

Electronic configuration = $2s^2 2p^3$

= 



≡ N ≡ N



σ_{P-P}

π bond

$\pi_{P_y - P_y}$,

$\pi_{P_z - P_z}$

In the case of P_x orbitals. They overlaps along Internuclear axis to give a sigma bond (P_x-P_x) then two, P_y orbitals, two, P_z orbitals overlaps laterally (sideways) overlapping to be two, π Bonds results. π-Bond is always a Weak bond. Because, the electron density spreads Perpendicular to the Internuclear axis. So that it is easily "exposed to the environment".

VB theory couldn't explain bond angles in NH_3 and in water molecule (H_2O). But it explained the bond angle as 90°.

* **Hybridization:-**

* (1) The process of mixing of atomic orbitals of same energy leads to formation of same numbered Hybrid orbital is known to be Hybridization.

(2) According to Hybridization paired valency orbital is also allowed for hybridization. But, these hybrid orbitals not allowed for chemical reaction.

(3) Hybrid orbitals with unpaired electrons are only allowed for Sigma bond (σ) formation.

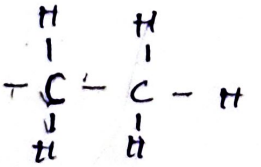
(4) To explain π-Bond formation 'unhybridized' Pure P-orbitals are allowed for π-bond formation.

(5) After the hybridization all hybrid orbitals are placed on central atom in such a way such that at which all the repulsions are minimised.

(6) These are sp^3, sp^2, sp, sp^3d, sp^3d^2, sp^3d^3 hybridization.

Examples:

(1) Ethane C_2H_6



In Ethane two carbon atoms are undergoes sp^3 Hybridization.

Atomic number of carbon = 6

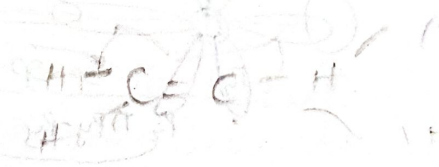
Ground state configuration = $2s^2 2p^2$

= [↑↓] [↑] [↑] [ ]

Excited state configuration = [↑] [↑] [↑] [↑]

According to sp^3 hybridization, the process of mixing of 1, 2s orbital, & 3, 2p orbitals leads to formation of 4, sp^3 hybrid orbitals. These hybrid orbitals on two carbon atoms arranged in such a way such that at which all the repulsions are minimised. 2, sp^3 hybrid orbitals of 2, carbon atoms overlaps along the inter nuclear axes to give a Sigma Known to be $\sigma_{sp^3-sp^3}$. The remaing 6, sp^3 hybrid orbitals of two carbon atoms overlaps with 6, 1S orbitals of six-Hydrogen atoms to give 6, σ_{sp^3-s}. The ideal geometry of the molecule is "Bi-tetra hydral" and the bond angle is "109°28'".

2) Ethelene :- C_2H_4



In Ethelene two carbon atoms are undergoes sp^2 Hybridization.

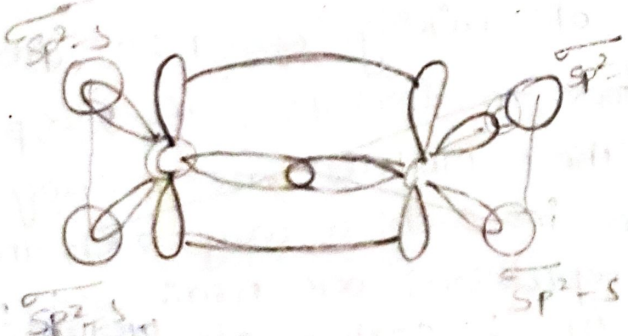
Atomic no. of carbon = 6

Ground state confi. = $2s^2 2p^2$

= [↑↓] [↑] [↑] [ ] ← vacant orbital.

Excited state confi. = [↑] [↑] [↑] [↑]

Total Bonds
↓
4
✓ σ ↓ 3 hybrid orbitals. sp^2
✗ π



Bi-triangular with Bond angle 120°,

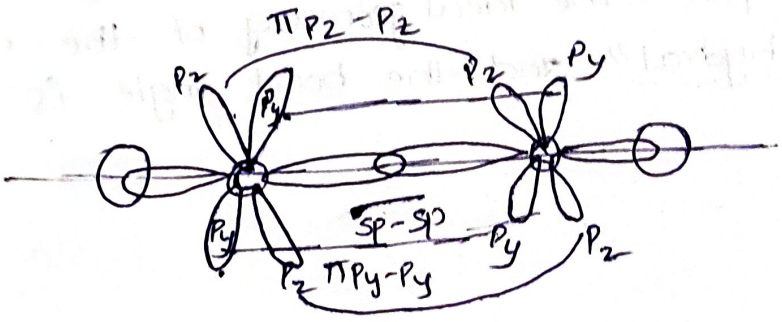
According to sp^2 Hybridization the process of

mixing of 1, 2S orbital and 2, 2P orbitals to give 3, sp^2 hybrid orbitals in which one sp^2 hybrid orbital of 1 carbon atom overlaps along internuclear axis towards another sp^2 hybrid orbital of carbon atom to give a sigma bond ($\sigma_{sp^2-sp^2}$). The remaing 4, sp^2 hybrid orbitals overlaps with the 4, 1S orbitals of 4, Hydrogen atoms to give 4, sigma bonds (σ_{sp^2-s}). The ideal geometry of the molecule is Bi-triangular with bond angle 120°

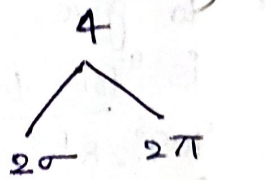
(4) AE ACETELENE : (C_2H_2)

(3)

$$H - C \equiv C - H$$



(Total Bond)

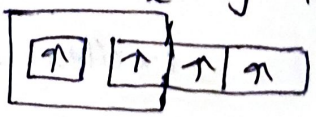


2H - SP - hybridiz
(2 Hybrid orbital

In Acetlene two carbon atoms ~~are~~ undergoes 'sp' hybridization.

Atomic number of carbon = 6

Excited state configuration = $2S^1 2P_x^1 2P_y^1 2P_z^1$



The process of mixing of 1, 2S orbital, 1, 2p orbital leads to formation of 2, SP hybrid orbitals. These orbitals are arranged on central atom in such a way such that at which all the repulsions are minimised. 2, SP hybrid orbitals of 2, carbon atom overlap along the internuclear to give a sigma bond (σ_{sp-sp}). The remaining 2, SP hybrid orbitals overlaps

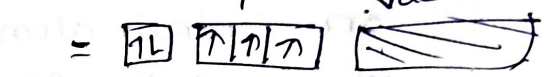
with 2, hydrogen atoms to give 2, sigma bonds (σ_{sp-s}). The remaining two unhybridized P-orbitals lateral overlaps to give two-π bonds ($\pi P_y - P_y$, $\pi P_z - P_z$). And the ideal geometry of the molecule is 180° (Linear)

(4) PCl_5 :- (Phosphorus Penta chloride)

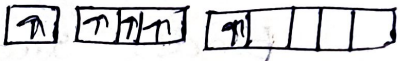
In PCl_5 the center atom phosphorus undergoes sp^3d hybridization.

Phosphorus atomic no = 15

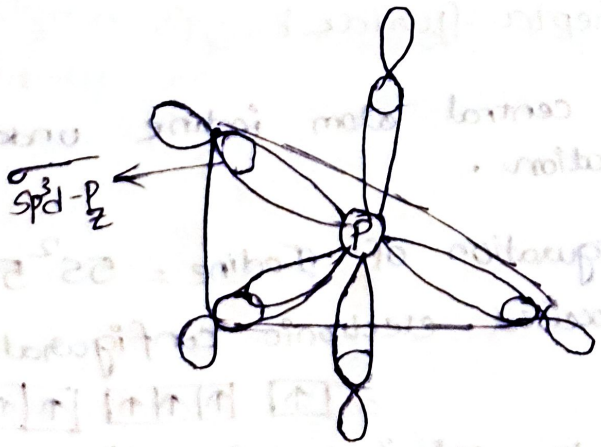
ground state configuration = $3s^2\,3p^3$



Excited state configuration = $3s^1 3p^3 3d^1$



According to sp^3d hybridization the process of mixing of 1, 3s & 3, 3p orbitals, 1, 3d orbitals to give 5, sp^3d Hybrid orbitals these are arranged on central atom at which repulsions are minimised. Then, all these 5, sp^3d orbitals (hybrid), overlaps with the 5, P_z orbitals of 5 chlorine atoms to give 5, $\sigma_{sp^3d-P_z}$. The ideal geometry of the molecule is "Trigonal bi-pyrimid".



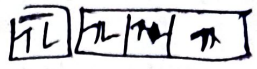
The ideal geometry of the molecule is Trigonal Bi-Pyrmid with the bond angles 120°, 180°.

(6) SF_6 :- Sulphur hexa chloride :

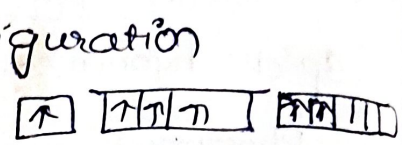
In SF_6 the central atom Sulphur undergoes sp^3d^2 hybridization.

Atomic no. of Sulphur = 16

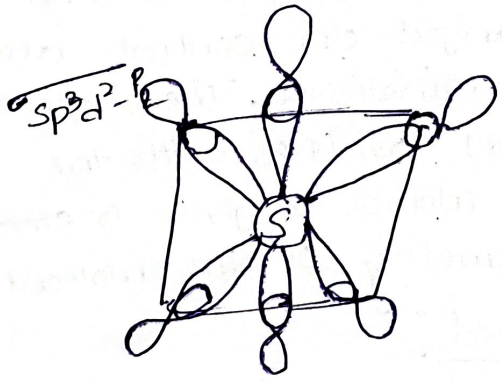
Ground State configuration = $3s^2 3p^4$

= 

Double excited state configuration

= 

According to sp^3d^2 hybridization 6, sp^3d^2 hybrid orbitals. These are arranged on central atom at which all repulsions get minimum. These 6 hybrid orbitals overlaps with 6, p_z orbitals of 6 Florine atoms to give 6, $sp^3d^2-p_z$ bonds. Ideal geometry of the molecule is 'Octahedral' and the bond angles are 90°, 180°.

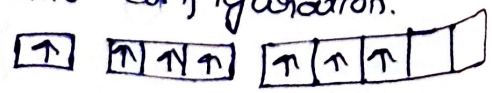


(7) IF_7 :- Iodine hepta fluride :

In IF_7 the central atom iodine undergoes sp^3d^3 hybridization.

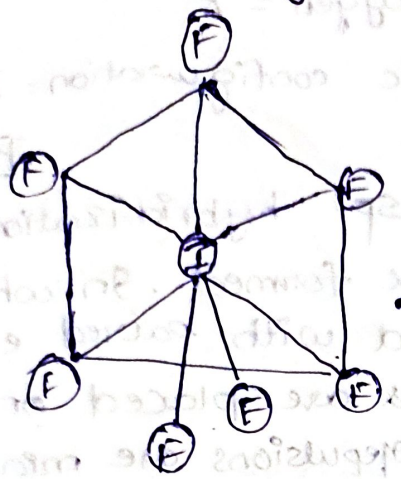
Electronic configuration of Iodine = $5s^2 5p^5$

The triple excited electronic configuration.

= 

According to sp^3d^3 hybridization, 7, sp^3d^3 hybrid orbitals are formed. These are arranged on central atom at which repulsions are

minimised. then, 7, P_z orbitals of 7, Florine atoms overlaps with the 7, sp^3d^3 hybrid orbitals to give 7, sigma bonds. (sp^3d^3-P_z). The ideal geometry of the molecule is "pentagonal bi-pyrmid" with the bond angles "110°, 180°."



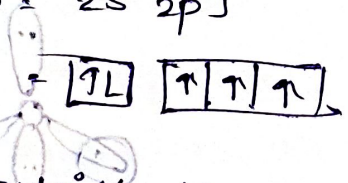
(8) NH_3:- (Ammonia)

1

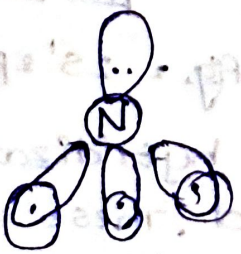
In ammonia the centeral atom Nitrogen undergoes sp^3 hybridization.

Atomic no. of Nitrogen = 7

Electronic configuration = $2s^2 2p^3$



According to sp^3 hybridization 4, sp^3 hybrid orbitals are formed. In which 1, sp^3 hybrid orbital is placed with 1, Loan pair of electrons. The remaining 3, sp^3 hybrid orbitals overlaps with the s-orbital of 3, hydrogen atoms to give 3, sigma ~~sp^3-s~~ (sp^3-s) bonds. The ideal geometry of the molecule is 'Pyramidal' with the bond angle 107°



(8) (9) & H_2O :-

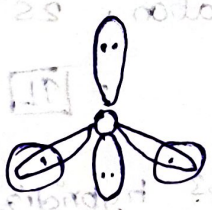
In H_2O the central atom oxygen undergo sp^3 hybridization.

Atomic no. of oxygen = 8.

Electronic configuration = $2s^2 2p^4$

= ↑↓ ↑↓ ↑ ↑

According to sp^3 hybridization 4, sp^3 hybrid orbitals are formed. In which 2, sp^3 orbitals are filled with paired electrons. All these orbitals are placed on central atom at which repulsions are minimised. 2, sp^3 hybrid orbitals overlaps with the S-orbital of two hydrogen atoms to give 2, (σ sp^3-s bonds). The ideal geometry of the molecule is angular (V-shaped) The bond angle is 104°30'.



104°30'

(10) Methane - CH_4

In methane the central atom carbon undergoes sp^3 hybridization.

Atomic no. of (C) = 6

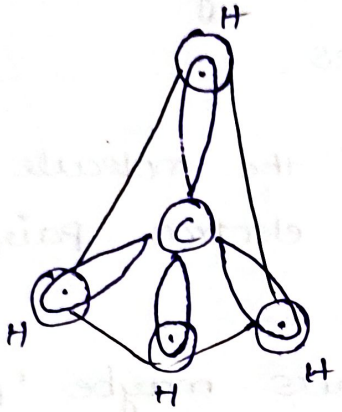
= $2s^2 2p^2$

Ground state = ↑↓ ↑ ↑

Excited state config. = $2s^1 2p^3$ ↑ ↑ ↑ ↑

According to sp^3 hybridization 4, sp^3 hybrid orbitals are formed. These are arranged at central atom at which repulsions are minimise These overlaps with the S-orbitals of 4, hydrog

atoms to give 4, (σ sp^3-s). The Ideal gemotry of the molecule is "Tetra hedral" and the bond angle is 109°28'.



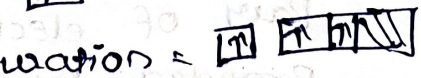
(b) BCl_3 :- (Boran tri-chloride)-

In BCl_3 the central atom Boron undergoes sp^2 hybridization.

Atomic number of Boron = 5

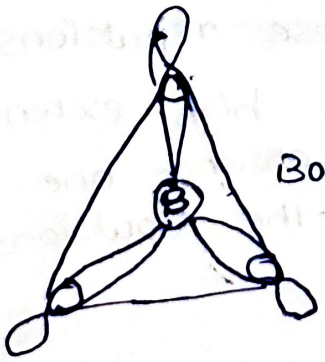
Electronic configuration = $2s^2 2p^1$

Grand state: 

Excited electronic configuration = 

According to sp^2 Hybridization 3, sp^2 hybrid orbitals are formed. These are overlaps with 3, p_z orbitals of 3, chlorine atoms to give 3, σ bonds. The Ideal geometry of the molecul is planar triangle with the bond angle 120°.

σ sp^2-p_z



Bond angle 120°

Essay

8M

Valency shell Electron pair repulsion theory VSEPR:-

1. According to this theory, electrons are considered as point charges.

2. The structure of the molecule mainly depends upon total no. of electron pairs on central atom.

3. These electron pairs maybe 'bond pairs & loan pairs.'

4. Bond pairs always placed between two centers of two atoms. where as, loan pairs always placed on "apex of central atom".

5. Since, bond pairs always placed between two atoms. They always occupy less space Provided by the molecule. where as, loan pair of electrons occupies much more space Provided by the molecule. which leads to the following repulsions in the molecule.

(i) Loan Pair - Bond Pair
(ii) Bond Pair - Bond Pair
(iii) Loan Pair - Loan Pair

6. Among all these repulsions loan pair - loan pair is having high extent. That's why the molecule itself arranges one particular structure at which all the repulsions are minimised.

SNO	Total no. of electron Pairs	Bond Pairs	lone Pairs	Structure	Bond angle.
01.	2	2	0	$BeCl_2$ linear	180°
02.	3	3	0	BCl_3 Triangular	120°
03.	4	4	0	CH_4 Tetrahedral	109°28'
04.	4	3	1	" NH_3 Pyramidal	107°
05.	4	2	2	H_2O Angular	104°30'

4m * *** Hydrogen bond :-

If hydrogen atom is placed between two highly electro-negitive elements by strong electrostatic attraction force then this type of attraction force is known to be hydrogen bond

Conditions for hydrogen bond:-

1. High electronegitivity.
2. Less atomic radii, are the conditions to show hydrogen bond.

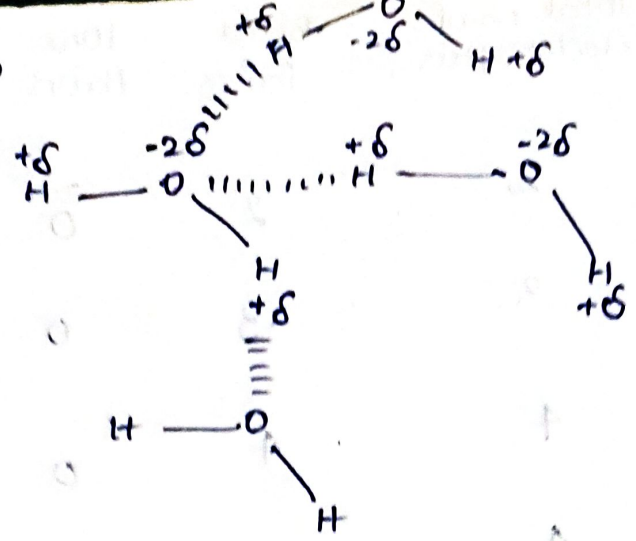
Hydrogen bond is 2 types:-

1. Inter hydrogen bonding
2. Intra hydrogen bonding

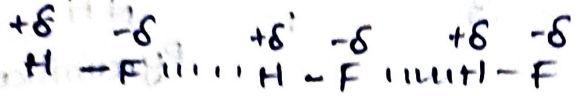
* Inter hydrogen bonding:-

If hydrogen bond is formed between similar molecules. Then, it is known to be Inter hydrogen bonding.

Eg:- H_2O



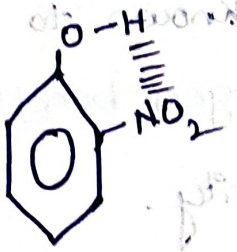
(ii) HF



(2) Intra hydrogen bonding:-

If hydrogen bond is formed between the molecule (within the molecule) is known to be Intra hydrogen bonding.

Ex:- Nitro phenol.



O-Nitrophenol

(Para)
(P-)
(ortho)
(O-)
(m-)
(meta)

Applications for Hydrogen bond:-

1. Due to the existance of Hydrogen bond the Boiling points of the liquids increases.

2. Due to the existance of Inter Hydrogen bonding (H_2O) is in liquid state where as, (H_2S) is in gaseous state.

REASON:- Since sulphur atom has got high atomic radius with low electronegitivity it doesn't form Inter hydrogen bonding

3. H_2O is having ~~having~~ high boiling Point (100°c) while 'HF' is 20°c.

REASON :- In the case of H_2O the extent of hydrogen bonds towards one water molecule is four (4). whereas, in the case of 'HF' the extent of hydrogen bonds are '2'. That's why H_2O has got high boiling Point.

4. HCl molecules doesn't form Inter hydrogen bond.

Reason:- Chlorine is having high electro-negitivity (3.1 units). But, due to high atomic radii. it is incapable to form inter hydrogen bonding.

5. 'NH_3' is having high boiling point.

Reason:- Due to less atomic radii, high electronegitivity (3.1 units). It can form Inter hydrogen bonding. that's why It is having Inter hydrogen bonding.

8M *** Molecular orbital theory :- (M.O. theory)

1. linear combination of atomic orbitals leads to formation of same numbered molecular orbitals (M.O's) in which half of the M.O's are having less energy than individual A.O's. Similarly, the remaining half of M.O's are having high energy than individual AO's.

A.O's - Atomic orbitals

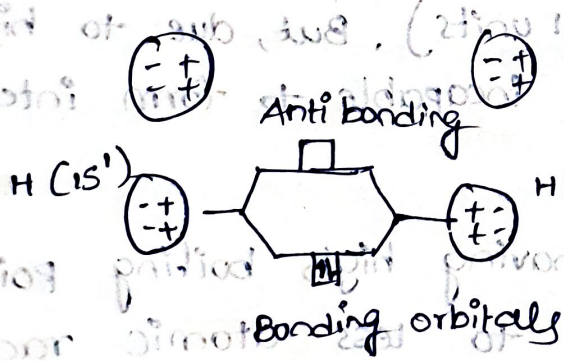
2. The lowest energy molecular orbitals are known to be bonding orbitals and high energy molecular orbitals known to be Anti bonding orbitals

3. According to M.O Theory bond order

Bond order $= \frac{1}{2}$ [Total no. of $e^{\ominus}$'s in Bonding orbital − Total no. of electrons in Anti bonding orbital]

MOED :- (Molecular orbital energy diagrams) :-

(1) H_2 Molecule :-



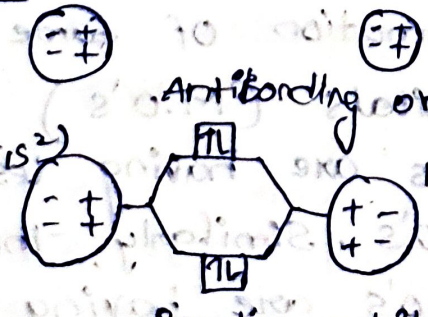
Bond order $= \frac{1}{2}(2-0)$

$= \frac{2}{2}$

B.O = 1

H − H

(2) He_2 Molecule :-



Bond order $= \frac{1}{2}(2-2)$

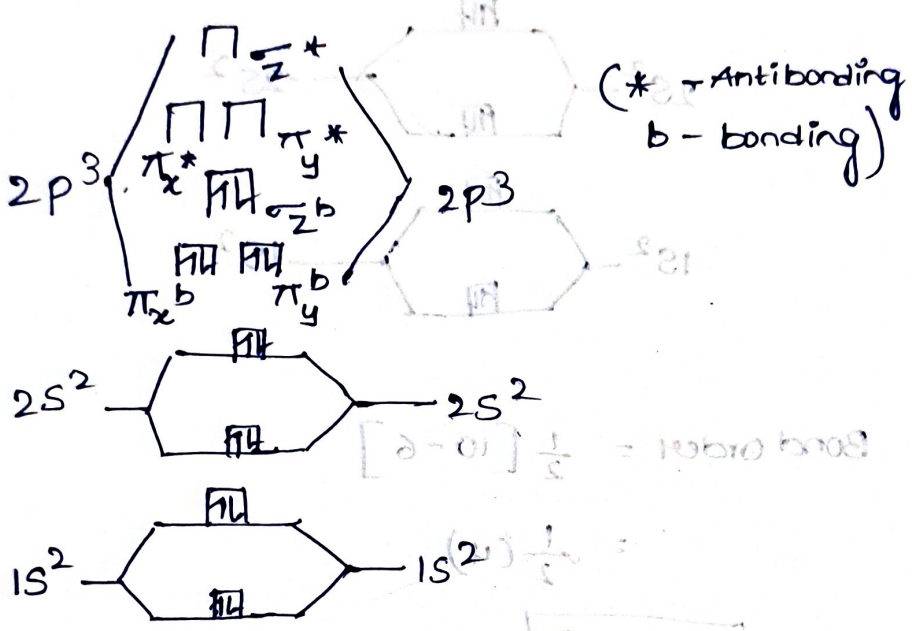
$= \frac{1}{2}(0)$

⇒ Bond order = $\frac{0}{2}$

$\therefore$ B.O = 0

According to M.O theory bond order of He_2 is '0'. That's why Helium never forms a chemical bond. because of completely filled configuration (Inert gas configuration).

(3) Atomic number = Z, $(Z \leq 7)$



(* - Antibonding
b - bonding)

here, Bond order = $\frac{1}{2}[10-4]$

$= \frac{6}{2}$

B.O = 3

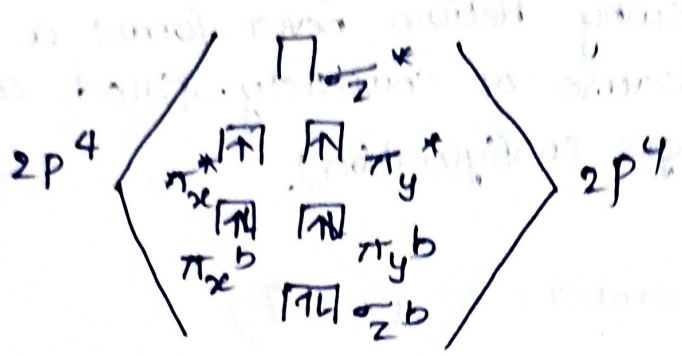
$\therefore$ N≡N

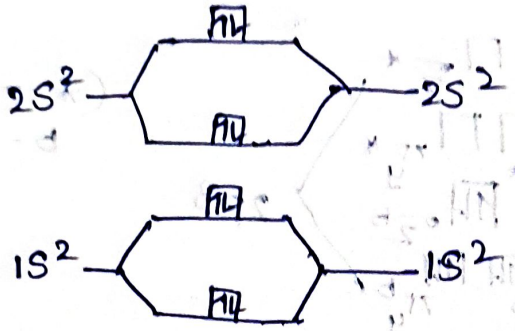
(4) Atomic

(4) Atomic number = Z, $Z \geq 8$, $Z \leq 20$;

Oxygen Molecule

O_2 :-





Bond order = $\frac{1}{2}[10-6]$

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

$\therefore B.O = 2$

$\therefore O = O$

The existance of two unpaired electrons in antibonding orbitals towards π-bond shows "Para-magnetic" behivious of oxygen molecule