—: Chemistry Exercise :—

Chapter no:- 03

~~Electrone~~

—: Chemical Bonding :—

QNO1 The Table shows the atomic Number and boiling Point of some noble gases.

Gas	Helium	Neon	Argon	Krypton	Xenon
Atomic no.	-12	-110	-118	-136	-154
Boiling Point C°	-253	-246	-186	-152	-107

(a) Explain this trend in Boiling Point

Boiling point ∝ Atomic Number

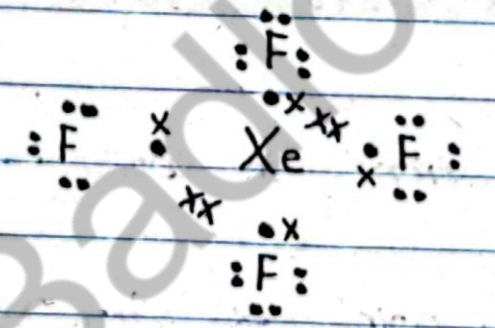
The Boiling point of noble gases increase with Atomic Number. Larger Atoms has more electrons. which enhances london dispersion force (van der waals force). As a Result, Stronger intermolecular forces aquires more energy to overcome, leading to Higher boiling point

Gas	Helium	Neon	Argon	Krypton	Xenon
Boiling Point K°	20°	27°	87°	120°	166°

(b) Xenon forms a number of covalently bonded compounds with fluorine.

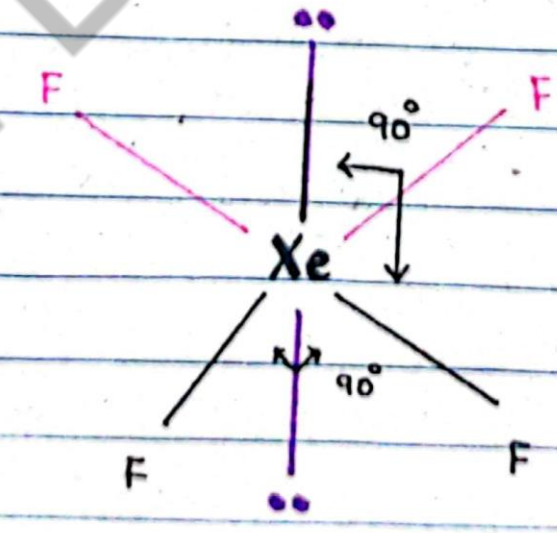
(i) Draw a dot-and-cross diagram for Xenon tetrafluoride XeF_4

Valence Electron of Xenon (Xe) = 8
Valence Electron of fluorine (F) = 7
Total Number of valence Electron in XeF_4 = 8 + (7 × 4) = 8 + 28 = 36



• Fluorine
× Xenon

(ii)

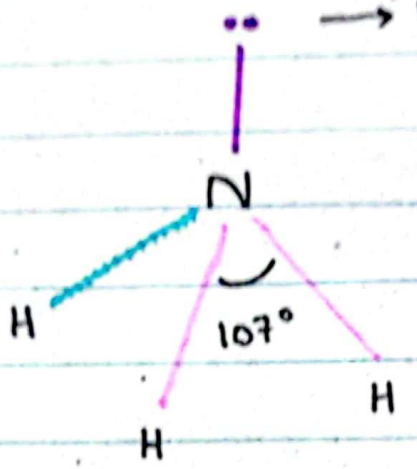


The Shape of XeF_4 is Square planar The central Xe atom has 4 bond pairs of electron and two lonepairs of Electron It undergoes sp^3d^2 has octahedral electron geometry and Square planar molecular geometry. The two lone pairs are at opposite corners of an octahedron. This minimize repulsion between them.

Q NO:-2 Aluminium chloride $AlCl_3$ and ammonia NH_3 are both covalent molecules.

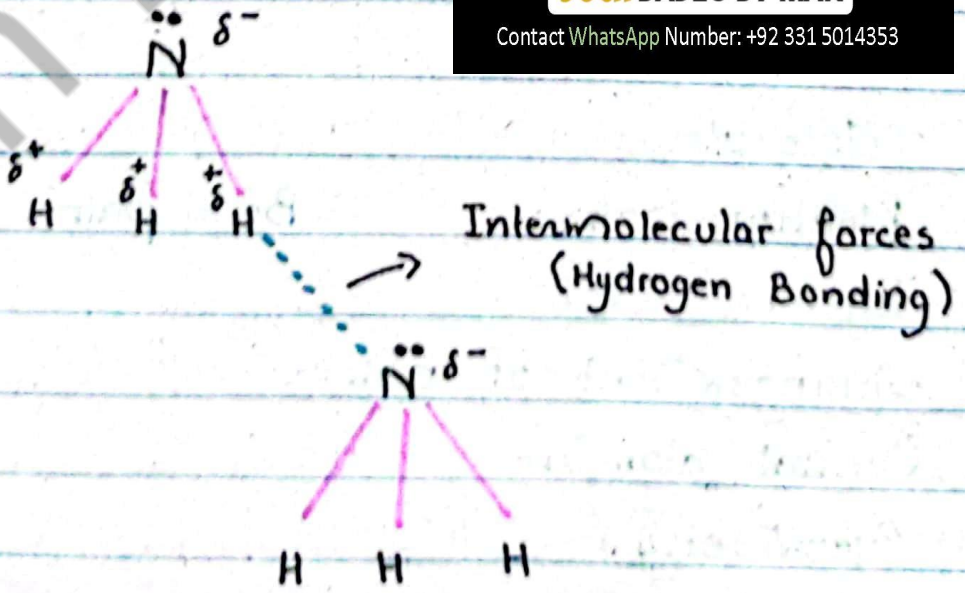
(a) Draw a diagram of an ammonia molecules Showing its Shape. Show any lone pair of electron. Also State the bond angle HNH in the ammonia molecules

Molecules :- NH_3 ammonia
Electron pairs :- 4 Bond pair :- 3 lonepair :- 1
Type :- AX_3E
Arrangement of pair :- Tetrahedral
Central atom :- $\overline{V}$ A
Geometry :- Trigonal Pyramidal
Bond Angle :- 107°



(b) What type of forces are present in ammonia molecules. Draw Diagram to show forces between ammonia molecules.

In Ammonia, Hydrogen Bonding is present between Nitrogen and Hydrogen. It is shown by dashes and it is stronger than london Dispersion or dipole-dipole forces.



(c) An Ammonia molecules and an aluminium chloride molecules can join together by forming co-ordinate bond.

(i) Explain how a Co-ordinate molecules is formed.

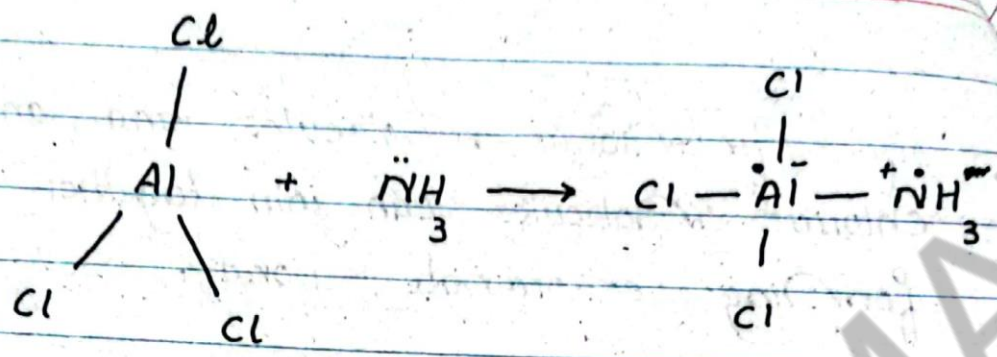
Co-ordinate Molecules:-

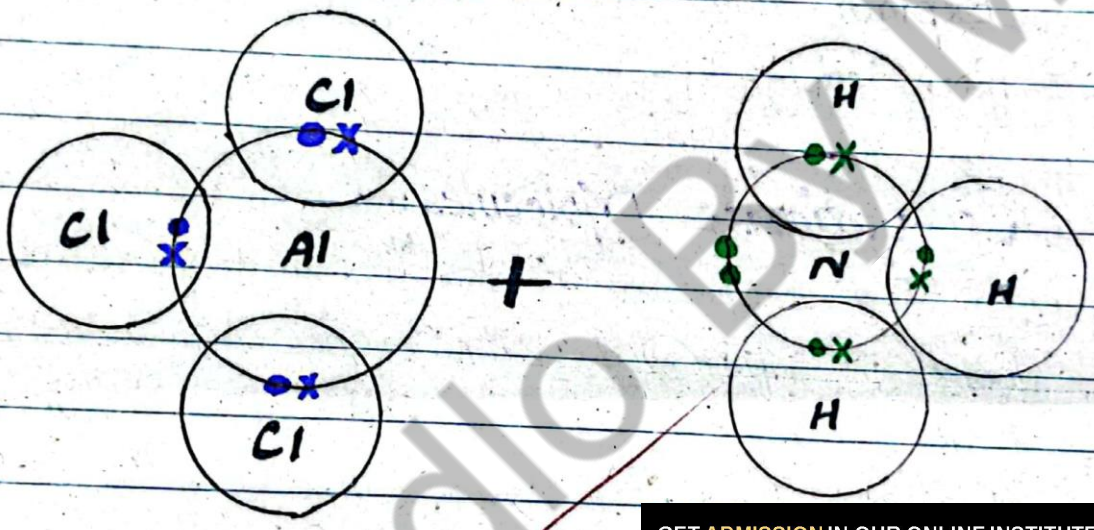
"Co-ordinate Covalent Bond is formed when one atom provides both the electron needed for covalent bond or Shared pair of Electron."

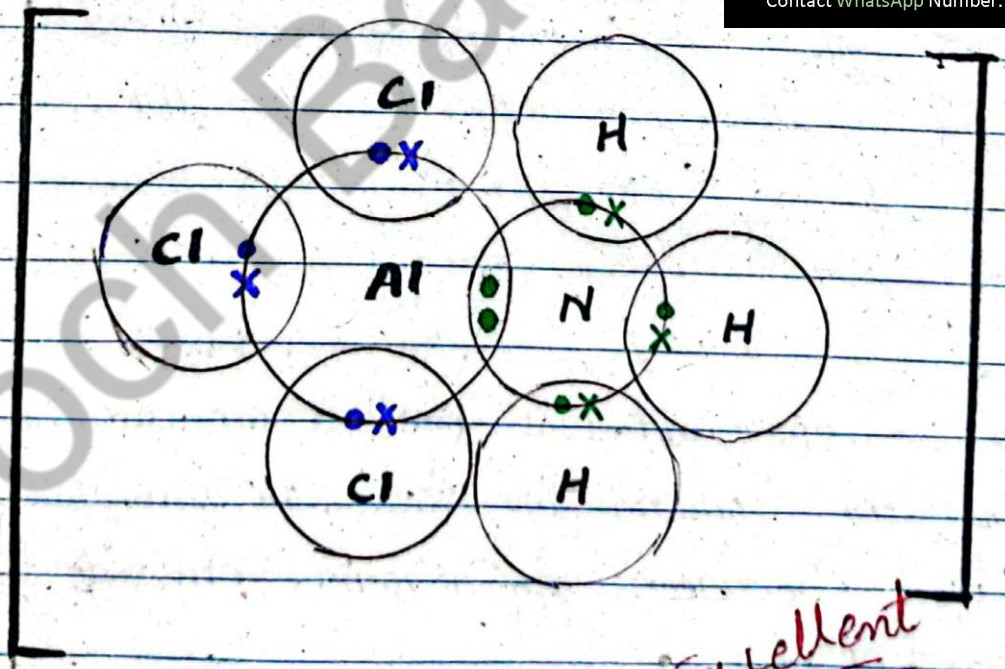
In Ammonia (NH_3) and Aluminium chloride $AlCl_3$, the Nitrogen atom in ammonia donates its lone pair electron to the empty orbital of aluminium atom. Creating co-ordinate bond.

(ii) Draw a dot-and-Cross ~~product~~ diagram to Show the bonding in the compound formed between Ammonia chloride

H_3NAlCl_3







Excellent

H_3NAlCl_3

Q NO:3 Electronegativity values can be used to predict the polarity of bonds.

(a) Explain the term Electronegativity.

Electronegativity :-

It is the power of covalently atom (Bonded atom) to attract Shared pair of electrons toward itself in a molecules.

It is a dimensionless property and has no unit Because it is only a tendency. It basically indicates the net result of the tendencies of atom in different elements to attract the bond-forming electron pairs.

(b) The Electronegativity values of some atoms are given below:

H = 2.1 , C = 2.5 , F = 4.0 , Cl = 3.0, I = 2.5

Use these values to predict the polarity of each of the following bond by coping the bonded atom Shown below and adding δ^+ or δ^- above each atom.

(i) $H - I$

Polarity = $H^{\delta+} - I^{\delta-}$

Δ Electronegativity = 0.4

Reason = In H-I molecules iodine being more electronegative (2.5) atom bear negative δ^- and hydrogen being less electro-negative atom (2.1) bears δ^+ charge.

(ii) $F - I$

Polarity = $F^{\delta-} - I^{\delta+}$

Δ Electronegativity = 4.0 - 2.5 => 1.5

Reason = In F-I molecules Fluorine being more electronegative atom bears δ^- (4.0). and iodine being less nelectronegative (2.5) bears δ^+ charge.

(iii) $C - Cl$

Polarity = $C^{\delta+} - Cl^{\delta-}$

Δ Electronegativity = 3.0 - 2.5 => 0.5

Reason = In C-Cl molecules carbon being less electronegative atom (32.5) bears δ^+

and Cl being more electronegative in nature (3.0) bears δ^- charge

(c) Describe the & Shape of ICl_3 molecules. Also mention its bond angle.

Molecules: ICl_3

Generic Formula (VSEPR):- AX_3E_2

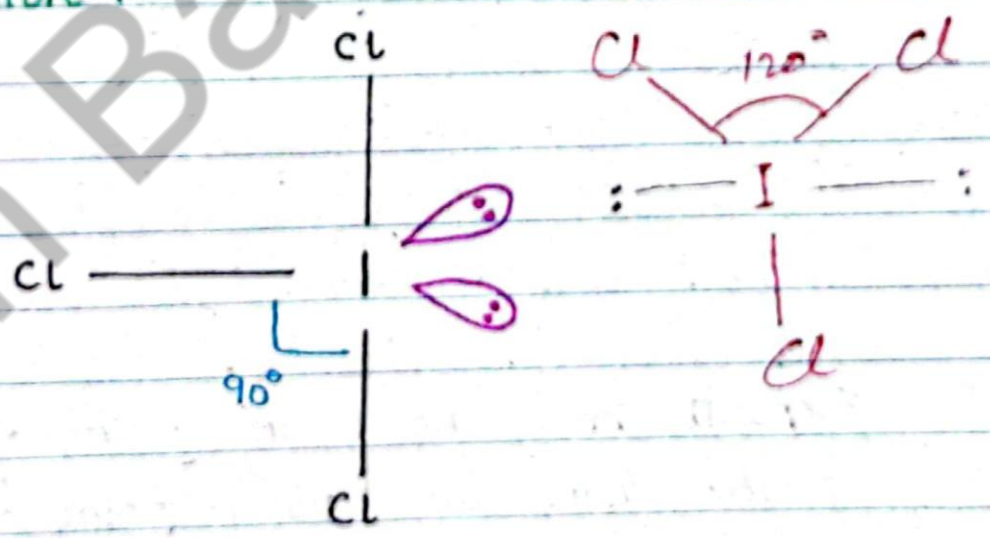
Bonded atoms: 03

lone pairs: 02

Molecular Shape: T-Shaped Trigonal

Bond Angle:- 90° 120° Planar

Structure:



Electron Geometry:- Trigonal bipyramid

d- The boiling point of the hydrogen halides are shown in the table

Hydrogen Halides :-	HF	HCl	HBr	HI
Boiling Point / °C :-	+20	-85	-67	-35

(i) Explain the tend in boiling Point from HCl to HI.

Hydrogen Halides	HF	HCl	HBr	HI
Boiling Point °K°	293	188	206	238

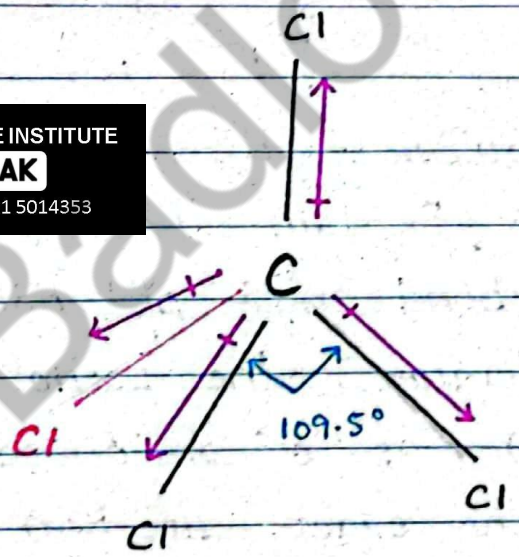
The boiling point from HCl to HI down the group is increase Because of the stronger london dispersion force therefore they require more energy to overcome Hence there boiling Point increase.

(ii) Explain why the boiling point of HF is greater than the boiling point of HCl.

This is because the Hydrogen bond in HF is stronger as compared to the weake

dipole - dipole interaction in HCl. Thus HF need more energy to break as compared to HCl. Hence HF has higher boiling point than HCl.

e- Tetrachloromethane CCl_4 is a non-polar molecules. Draw a Diagram to show the shape of this molecules. Explain why this molecules is non-polar.

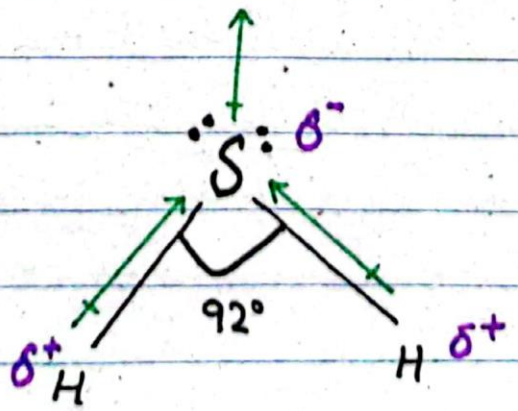


Tetrachloromethane (CCl_4) has a tetrahedral shape with the carbon atom at the centre and four chlorine atoms at the corner. The bond angle is 109·5°

Reason for Non-Polarity :-

Carbon tetrachloride (CCl_4) is nonpolar because of its symmetrical tetrahedral molecular geometry. Although the C–Cl bond are polar due to the difference in electronegativity between carbon and chlorine, the molecule's symmetrical shape causes the dipole moments to cancel each other out. As a result $CICl_4$ has no net dipole moment, making it non-polar overall.

Q4 Hydrogen Sulphide H_2S, is a covalent compound. Explain the type of Hybridization also write bond angle of HSH. Also show on your diagram the partial charges on each atom as δ^+ or δ^- and an arrow showing the exact direction of the dipole in the molecules as a whole.



• Direction of overall dipole

$\mu = 0.95D$

Hybridization:-

In Hydrogen Sulphide (H_2S) has sp^3 hydridized. This means that Sulfer undergoes hybridization to form Two Single bond and 2 lone pair bond. While forming a bond the s orbital of the Hydrog atom overlaps with sp^3 hybrid orbital of a Sulfer atom.

Bond Angle:- The bond angle in HSH is 92° Degree. Slightly less than the ideal tetrahedral angle of 109.5° Degree. This is due to the repulsion from lone pairs.

Dipole Moment:- Sulpher is more electronegativ than Hydrogen So, Sulpher has a partial negative charge (δ^-) and hydrogen has a partia positive charge (δ^+). The direction of dipole moment is from Hydrogen to Sulpher ($H \rightarrow S$).

(b) Explain why the ~~boiling~~ hydrogen Selenide, H_2Se, has a higher boiling point than Sulphede, H_2S

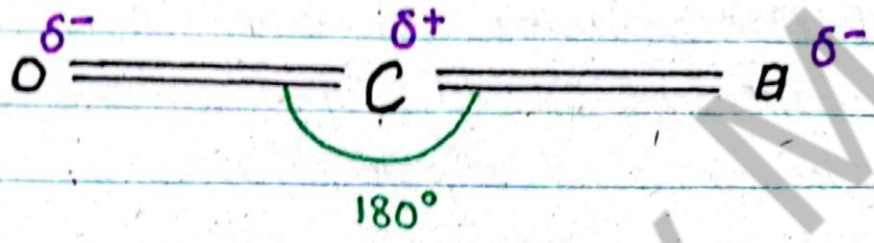
Hydrogen Selenide (H_2Se) has a higher boiling point than Hydrogen Sulphide (H_2S) Because of the intermolecular forces in H_2Se Because Selenium is larger than Sulphur As a Result Stronger London dispersion forces in it and harder to break it. These Stronger dispersion forces requires more energy to overcome. Hence its boiling point goes higher from H_2S.

) Explain why the boiling point of water is so much higher than the boiling Point of Hydrogen Sulphide.

The boiling point of water is much higher than that of hydrogen Sulphide because water form a Stronger hydrogen bond between its molecules. Hydrogen bonding is much stronger than dipole-dipole interaction in H_2S. Hence due to stronger bonding it is need more energy to break this bond As a Result its boiling point is Higher than H_2S.

Q NO 5

(a) Describe the shape carbon dioxide molecules.



- It has a linear shape.
- Its bond angle is 180° Because it form 2 bond bo double bond pair
- All three atoms has completed their octet. So, this molecule is stable.
- Carbon has less electronegativity than O so it is Slightly δ^+ and Oxygen is Slightly δ^-
- This net dipole moment is zero.

(b) Bromine is liquid at room temperature Weak van der Waals' forces hold the bromine molecules together. Describe how van der Waal's forces arise.

Van-der-Waals forces arise from temporary dipole-~~dipole~~ induced dipole forces. When molecules are close proximity, the uneven election within eachmolecules can cause temporary induced dipole. This dipole interact with induced dipole moment of neigboring molecules creating a weak attraction force known as "van-der-Waals force"

The Strength of van-derWaals depend upon Polarity of molecules. like bromine being High Polarizable molecules experience Strong van-der-Waals forces as compared to less polarizable molecules.

Q.No-7 Water is extensively hydrogen bonded. This gives it anomalous (peculiar) Properties.

(a) Explain why ice is less denser than liquid water. State two other anomalous Properties of water.

Ice is less Denser than Water :—

Ice is a rigid structure with molecules spaced farther apart due to hydrogen bonding. Making it less Denser than water where molecules are closer together.

Anomalous properties of Water :—

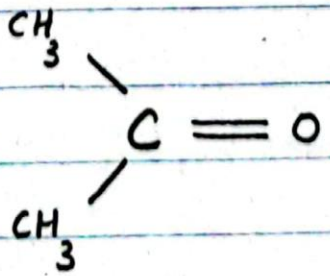
1- **High Melting and boiling point :-**

Water boils at 100°c and melts at 0° because of Stronger Hydrogen bond.

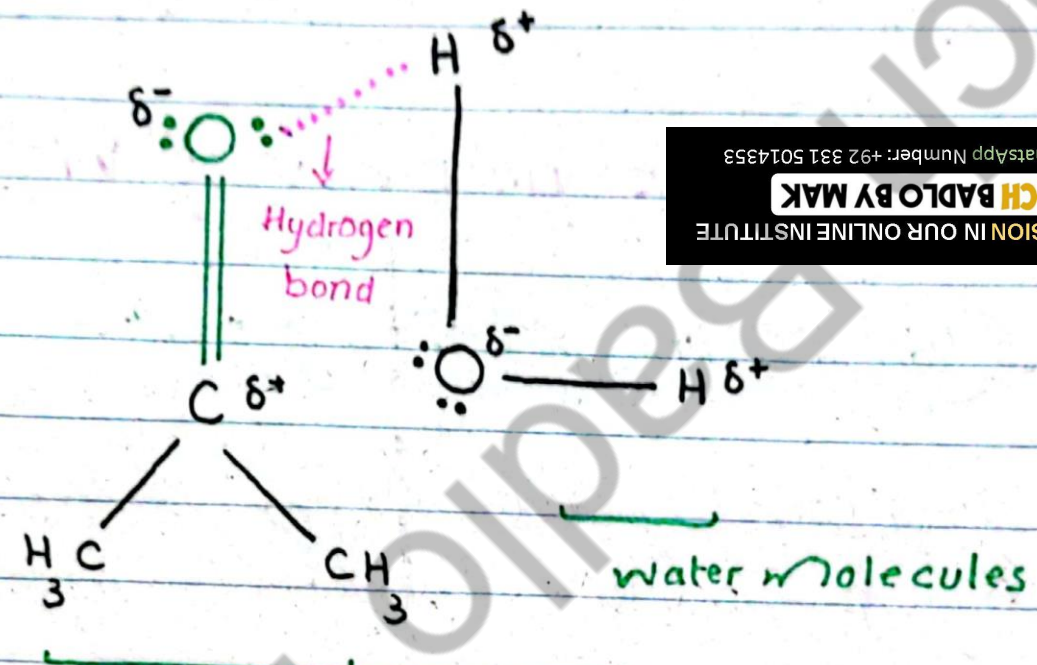
2- **High Specific Heat Capacity :-**

Water absorbs and release much heat without changing temperature because of Hydrogen bonding.

(b) Propanone has the structure shown below.



(c) (i) When propane dissolve in water, it forms a hydrogen bond with water. Draw a diagram to show propane molecules and a water molecules forming a hydrogen bond.



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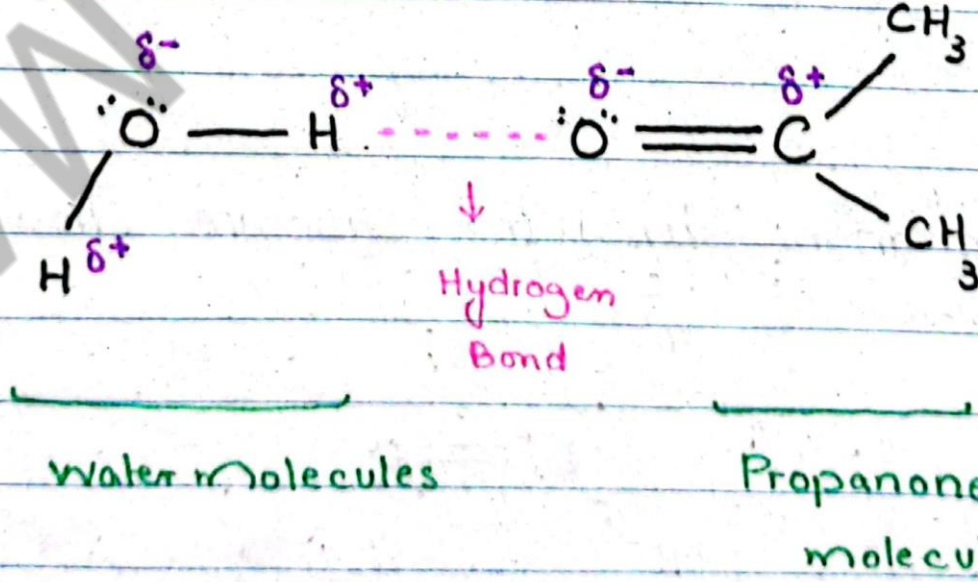
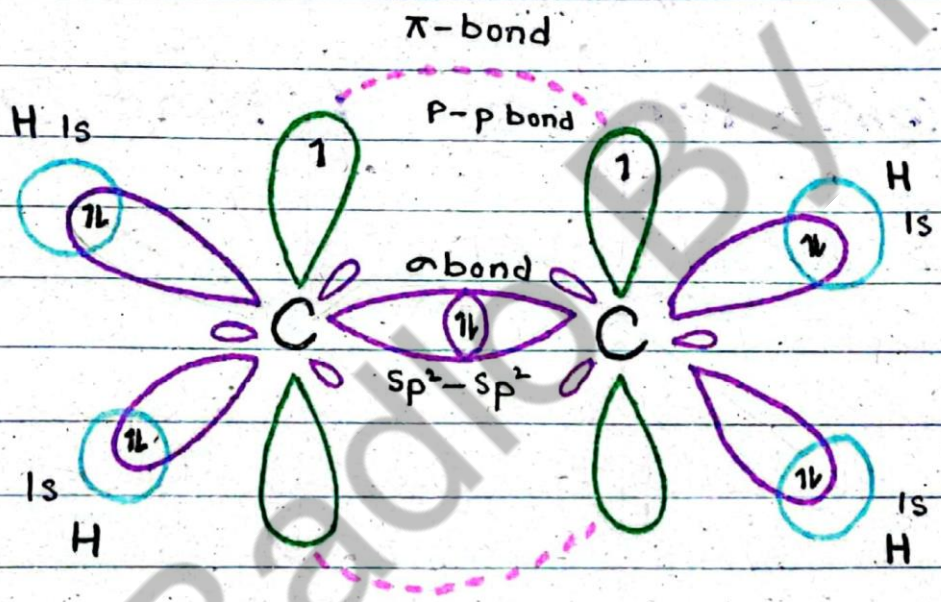


(C) (Also same as QNo1-11)

(i) Propanone has a double bond. One of the bond is σ bond (Sigma bond). The other is a π (pi bond). Explain the difference between σ bond and π bond in terms of how they are formed.

	Sigma bond (σ)	Pi bond (π)
1-	It is formed by head-on overlap of orbital	It is formed by side-by side overlap of parralle orbital.
2-	On the Bond axis ē density.	ē density above or below the bond axis.
3-	~~Weaker Bond~~ Stronger Bond.	~~Stre~~ Weaker bond
4-	Rotation of either atom doesn't break the bond	Rotation of either atom break the bond
5-	Present in Single double, triple bond	Present in double and triple bond.
6-	Bond is formed by hybrid orbit.	Bond is formed by unhybrization.

(ii) Copy the diagram, then complete it to show the shape of the electron clouds in the σ bond and the π bond between the carbon atoms in ethene. label your diagram.



Q No. 8 Energies of orbital can be explained by molecular orbital theory. It has been observed that in case of Nitrogen molecules $\sigma 2p_x$ is higher in energy than $\pi 2p_y$ and $\pi 2p_z$.

(a) Draw molecular orbital energy diagram for nitrogen molecules.

Tips
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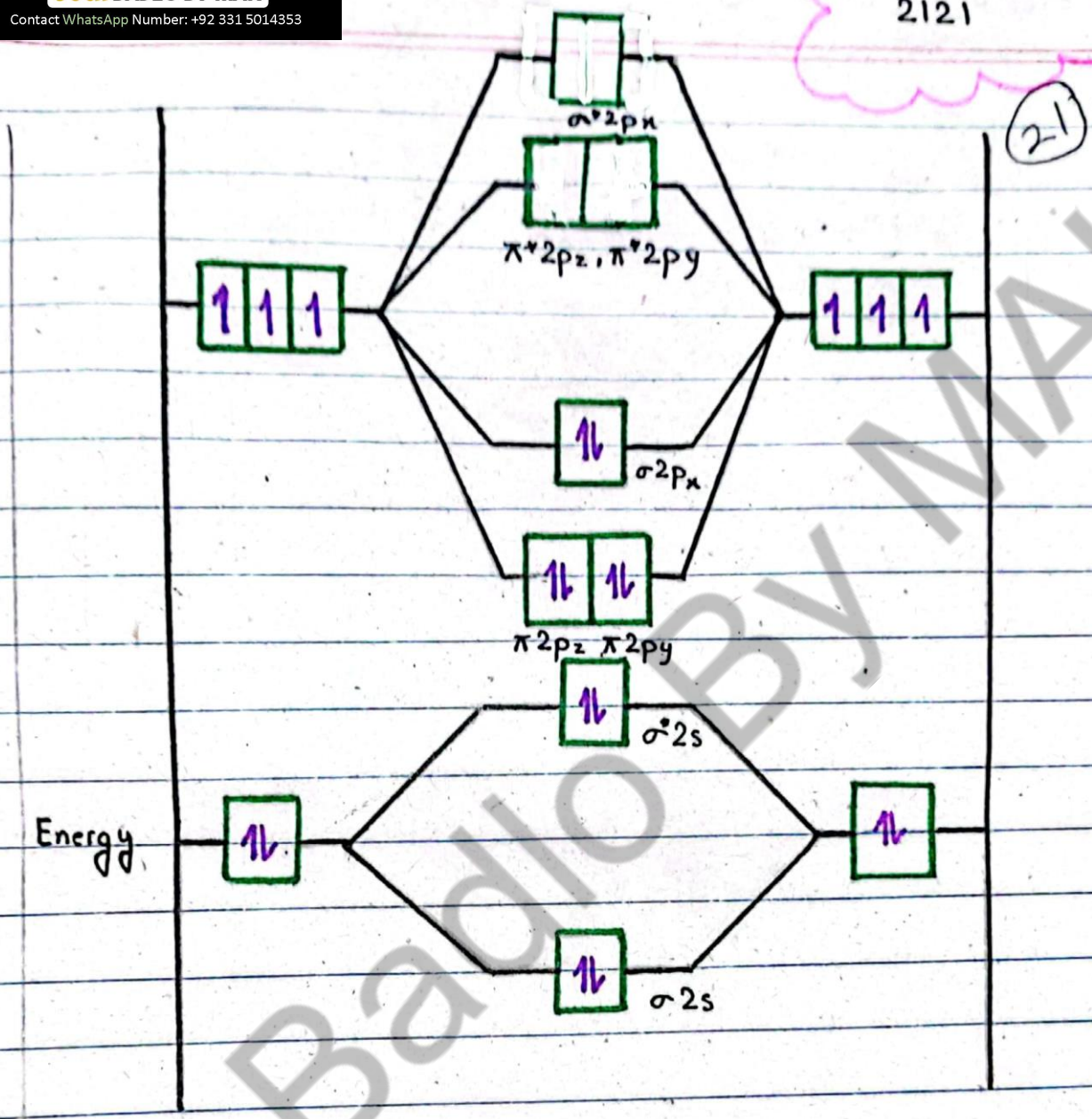


Nitrogen contains three unpaired electron in 2p orbitals.

Electronic Configuration :- $1s^2, 2s^2, 2p_x^1 2p_y^1 2p_z^1$

Total number of electron in N_2 molecules = 14

Bond Order = $\frac{18-2}{2} = \frac{6}{2} = 3$

(b) Give reason why the $\sigma 2p_x$ energy is greater than $\pi 2p_y$ and $\pi 2p_z$.

In a Nitrogen molecules (N_2), the $\sigma 2p_x$ orbital is higher in energy than $\pi 2p_y$ and $\pi 2p_z$ because

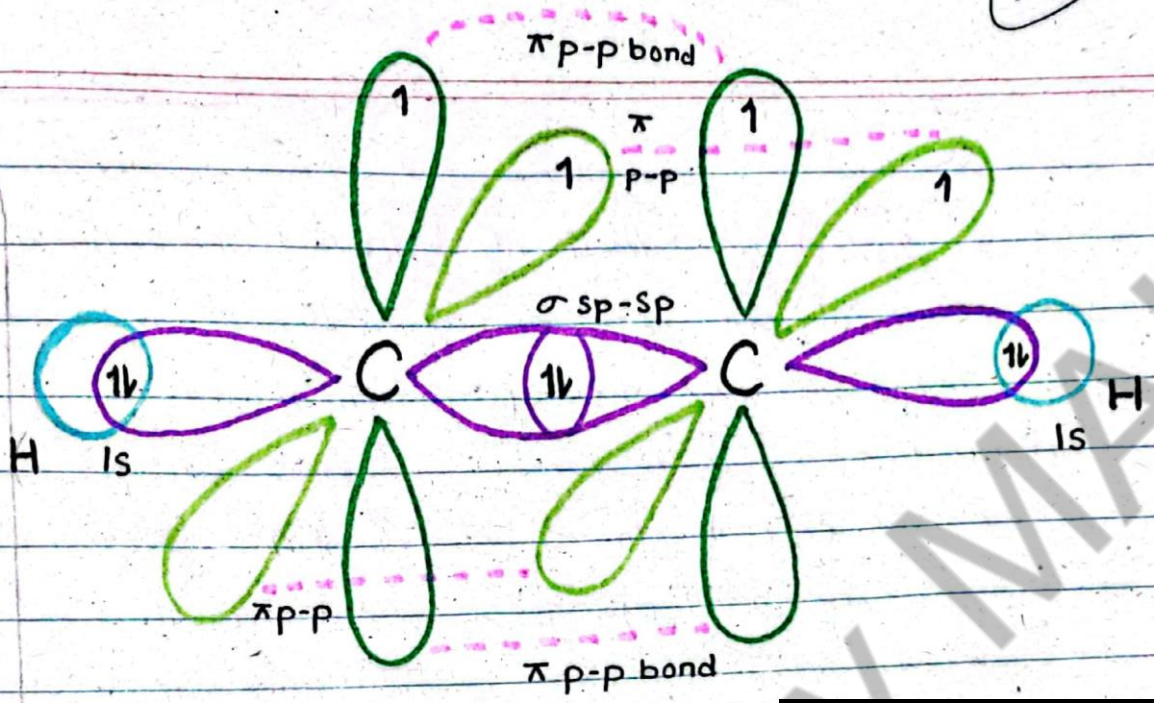
$\sigma 2p_x$:-

In $\sigma 2p_x$ the σ orbital is formed from end to end overlap which is stronger and ~~need~~ have higher energy.

$\pi 2p_y$, $\pi 2p_z$:-

In $\pi 2p_y$ and $\pi 2p_z$ the π orbital is formed from side by side overlap which is weaker bonding and have less energy than $\sigma 2p_x$. See book

Q.No-9 Carbon can make a bond with hydrogen to form ethyne : Bond energy of C–H is same although 2s and 2p orbitals are involved which have difference in energies. Explain the formation of ethyne molecules on the basis of hybridization with the help of diagram.



Ethyne Molecules (C_2H_2) :-

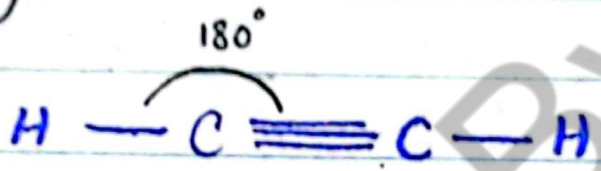
In Ethyne (C_2H_2) the central atom is Carbon. In groud ground state electronic Configuration of 'C' is $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z$ There are only two unpaired electron but the valency of carbon is 4 so it make four bond. So, the Configuration of 'C' in excited state is $1s^2\ 2s^1\ 2p_x^1\ 2p_y^1\ 2p_z^1$

Process of Hybridization :-

Each carbon atom undergoes 'sp' hybridization by using 2s and one 2p orbital which are arranged linearly. The two carbon atoms form a σ

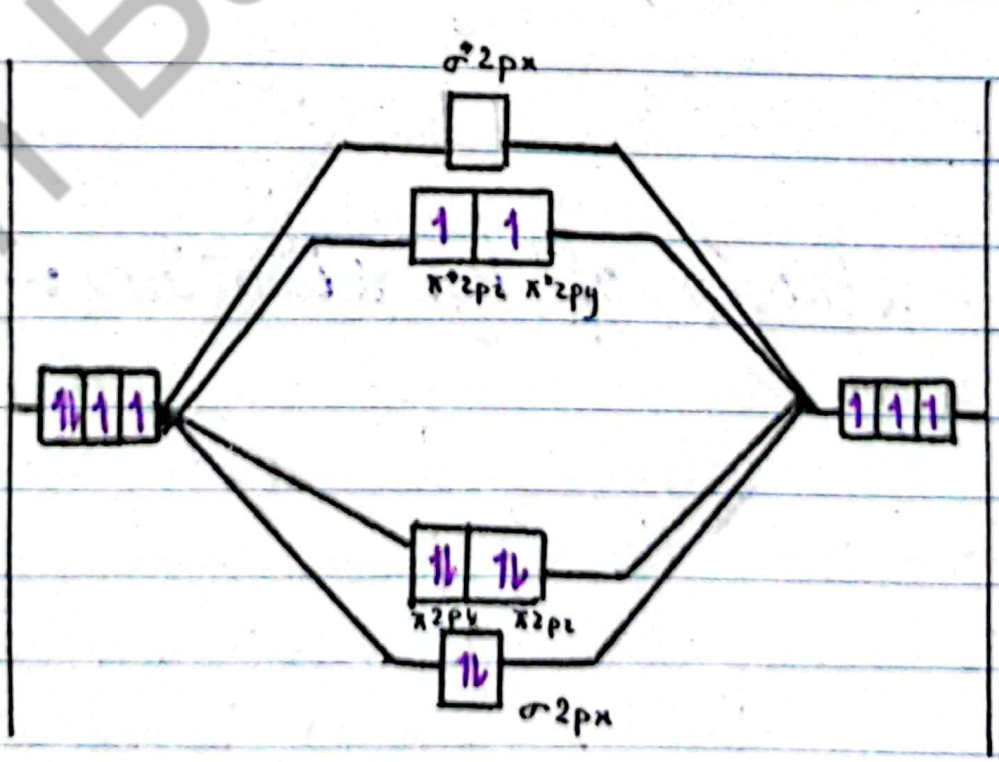
Sp-Sp bond. There are also two unhybrized p orbitals i.e, 2py and 2pz on each carbon atom which are peripendicular to sp hybrid orbitals.

Each carbon atom also formed sigma bond with H by overlapping of sp-s orbitals Its Bond angle is 180° and have linear Geometry.



Q No:-10 Explain the Magnetic properties of O_2, O_2^+ and N_2 by applying Molecular orbital theory.

O_2



Electronic Configuration of O_2 :- $1s^2 2s^2 2p^4$

Molecular orbital Electronic Configuration

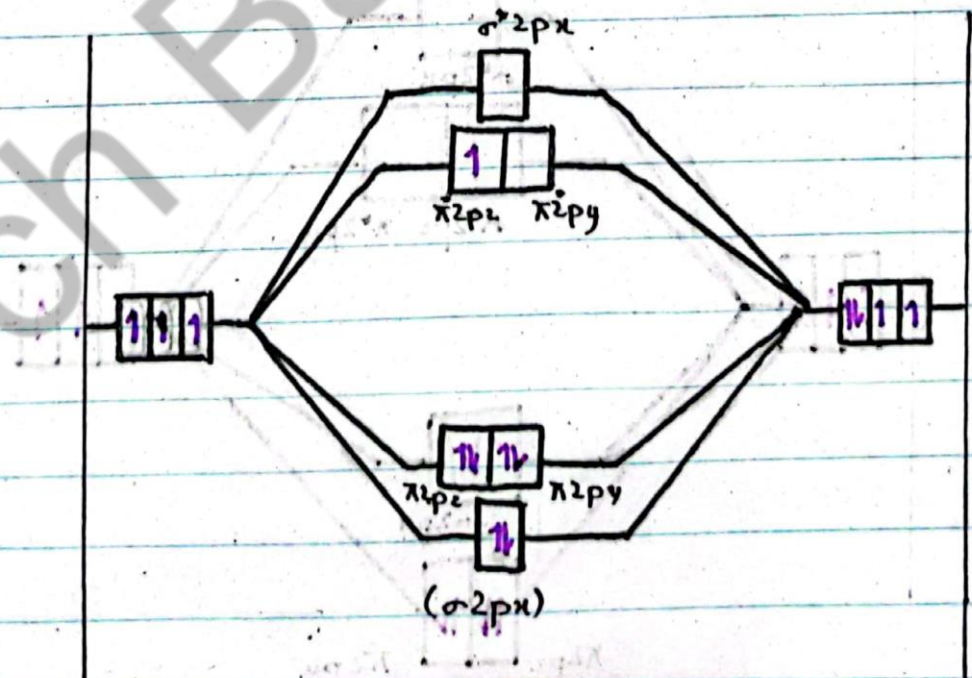
$(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_x)^2 (\pi 2p_y)^2 = (\pi 2p_z)^2 (\pi^* 2p_y)^1 = (\pi^* 2p_z)^1$

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

Magnetic Properties :-

Because each oxygen O has two unpaired electrons in molecular orbitals so it is paramagnetic in nature.

O_2^+



Electronic Configuration of O^+: $1s^2 2s^2 2p^3$

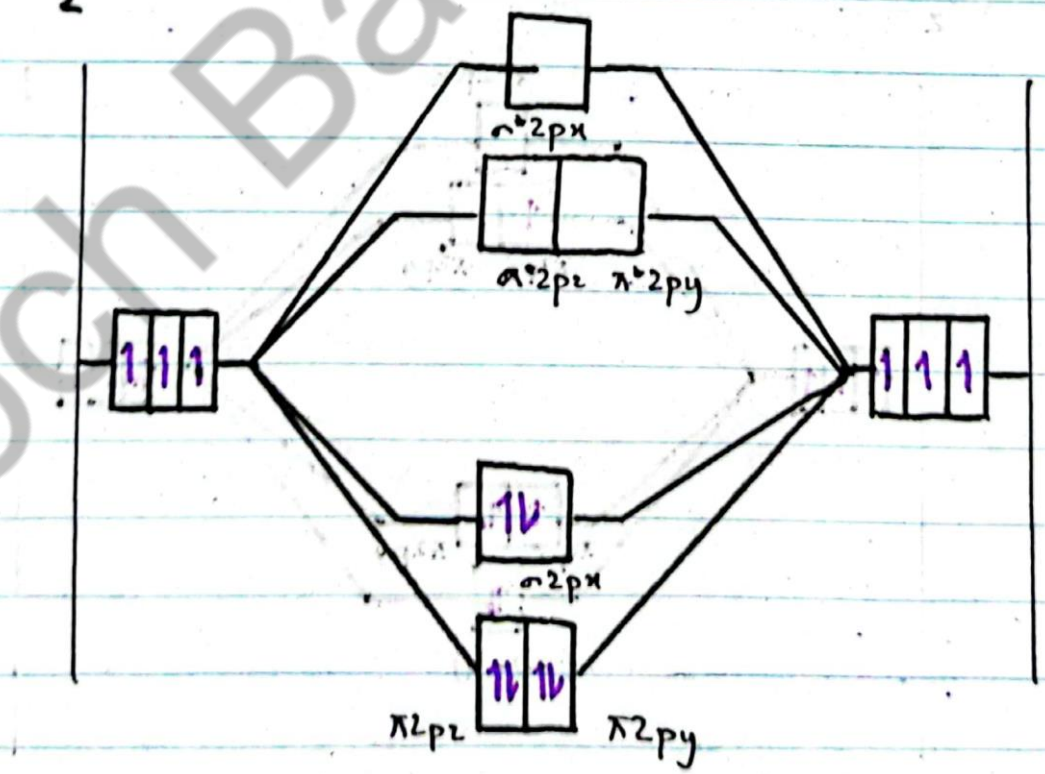
Molecular Orbital electronic Configuration O_2^+:
$(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\pi 2p_y)^2 (\sigma 2p_x)^2$
$(\pi 2p_y)^2 = (\pi 2p_z)^2 (\pi^* 2p_y)^1$

$$\text{Bond Order} = \frac{10-5}{2} = \frac{5}{2} = 2.5$$

Magnetic Properties

O_2^+ has one unpaired electron in Molecular Orbital so it is paramagnetic in nature.

N_2



Electronic Configuration of N = $1s^2 2s^2 2p^3$

Molecular Orbital electronic Configuration of N_2 = $(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\pi (2p_y)^2 = (\pi 2p_z)^2 (\sigma 2p_x)^2$

Bond Order = $\frac{10-4}{2} = 3$

Magnetic Properties :—

In N_2 there is no unpaired electron in Molecular Orbital so it is diamagnetic in nature.