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Class 11
Chemistry

Chapter 3
Chemical Bonding
Handwritten Notes

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28th August, 2025 Day: Thursday

CHAPTER: 3

CHEMICAL BONDING

ELECTRONEGATIVITY

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DEFINITION:

"It is the power of a covalently bonded atom to attract shared pair of electrons towards itself in a molecule."

Properties:

1- It is a dimensionless property and has no units because it is only a tendency.
2- It indicates the net result of the tendencies of atoms in different elements to attract the bond-forming electron pairs.

Instrument used:

Electronegativity is measured on several scales. The most commonly used scale is "pauling scale". This scale was designed by "Linus Pauling".

∴ The higest electronegativity of the element in periodic table is Flourine (F) that is 4.0 and the lowest electronegativity of the element in periodic table is Cesium (Cs) and Francuim (Fr) that is 0.7.

FACTORS AFFECTING ELECTRONEGATIVITY:

It depends upon.

- Size of atom
- Nuclear charge
- Screening by inner electrons.

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- As the number of proton in nucleus increases, the nuclear charge increases. Hence, the size of atom decreases causing the attraction of bonded electrons with the nucleus to decreases; resulting in <u>increase</u> of electronegativity

- When the number of inner shell electrons increases, the distance between the nucleus and valence electrons increases so, attraction of bonded electron with the nucleus decreases. Causing <u>decrase</u> in electronegativity

- $E.N \propto \frac{1}{\text{atomic size}}$
- $E.N \propto \frac{1}{\text{shielding effect}}$
- $E.N \propto$ nuclear charge

Trends In Periodic table:

Down the group:

Electronegativity decreases from top to bottom in group.

Reason:

- There is an increase in atomic number.
- The nuclear charge increases but the effect of the increase in nuclear charge overcome due to increase in shielding effect.
- The attraction of shared electrons with the nucleus decreases.

Example:

In the halogen group as we move down from fluorine to astatine the electronegativity value

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decreases.

F > Cl > Br > I > At

Along the period:

Electronegativity increases from left to right in period.

Reason:

- Effective nuclear charge increases.
- Shielding effect remains constant.
- Size of atom decreases.
- The power of atom to attract shared pair of electron increases.

Examples:

In the third period from sodium to chlorine the electronegativity value increases.

Na < Mg < Al < Si < P < S < Cl

* Electronegativity of first group (IA) is the lowest and the 7th group (VII) is the highest. Nobel gas do not have highest E.N because their outermost shell is completely fill and they do not gain e⁻s

Nature of Bonds in terms of Pauling Electronegativity values:

1. In ionic bonds, the difference in electronegativity between two atom higher than 1.8. (ΔE.N > 1.8)
2. Polar covalent bonds have difference in electronegativity b/w 1.8 and 0.4. (ΔEN = 1.8 to 0.4)
3. Non-polar covalent bond have differences in E.N of 0.4 or lower. (ΔE.N = 0 to 0.4)

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4- In case the electronegativity of both atoms is the same the pair of electrons will be shared equally.

For e.g:- $\overset{\delta+}{H}-*-\overset{\delta-}{H}$

2.1 2.1 $\Rightarrow 2.1-2.1=0$

Example:3.1 (Page:43)

1- **Oxygen molecule (O=O):**

This molecule consists of two oxygen atoms. Oxygen has an E.N value of 3.44. So each atom has E.N of 3.44. There difference in E.N is 0. which means none of the two atoms in the molecule pulls the electron to itself more strongly than the other. So the molecule is covalent bond.

O=O $\Rightarrow$ E.N = 3.44 - 3.44

3.44 3.44 = 0

2- Potassium Chloride (KCl):

This compound has potassium atom with an E.N value of 0.82 and a chlorine atom with an E.N value of 3.16.

K-Cl $\Rightarrow$ E.N = 3.16 - 0.82

0.82 3.16 = 2.34

The difference of E.N is greater than 1.8 so, the bond is ionic bond.

3- Hydrogen Chloride (HCl):

The molecule consists of hydrogen atom with an E.N of 2.0 and chlorine with an E.N of 3.16.

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H–Cl ⇒ E.N = 3.16 - 2.20
2.20 3.16 = 0.96

Chlorine is more E.N than hydrogen. this means Cl pulls the electrons more closely to itself than the hydrogen atom. The E.N value is 0.96 which is between 1.8 and 0.4. So the bond is polar covalent bond.

Covalent Character in a Compound:

"Covalent character in a molecule depends upon the polarising power of cation which in turn depends on the Oxidation State of cation."

Along the Period:

When we move from left to right in period covalent character increases.

Reason:

As the polarising power of cation is higher more will be the covalent character.

Example:

Aluminium Chloride ($AlCl_3$) is more covalent in nature as compared to Magnesium Chloride ($MgCl_2$). Oxidation state of aluminium atom is +3 and Magnesium is +2. As the polarising power of Al(+3) is higher it can polarise the anion causing electrons to be shared between two ions. Hence, covalent character will be developed in the compound.

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DIPOLE MOMENT

DEFINITION:

"It is the product of the magnitude of positive or negative charges and the distance between them."

MATHMETICAL FORM:

$$\mu = q \cdot r$$

dipole moment : Charge · seperation distance

where $q = 1.602 \times 10^{-19} C$

* dipole moment is a vector quantity

SI Unit:

The SI unit of dipole moment is debye.

→ 1 Debye = 3.335×10^{-30} Cm

→ Cm is coulomb meter.

Electronegativity Difference:

E.N difference between two atoms determines the bond polarity of the molecule. greater the E.N difference b/w two atoms, more will be the bond polar (more ionic nature). but a small difference in E.N of two bonded atoms will make bond less polar or covalent in nature.

$\mu = 0$ → (non polar bond)

$\mu \neq 0$ → (polar bond)

* The difference between the E.N of two atoms in a compound determines the overall dipole moment and overall polarity of the compound.

$H^{\delta+} - Cl^{\delta-}$

dipole moment has magnitude and directions

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POLAR AND NONPOLAR COVALENT BOND

DEFINITION:

"In polar covalent bond, due to electronegativity difference, one atom which is more electro-negative will attract shared electron to itself and bear partial negative charge and other bears partial positive charge. Such molecules are called polar molecules."

* Polar molecules have dipoles, electric charges of equal magnitude and opposite sign.

Non Polar Covalent Bond:

"In these molecules dipole of bond exerts equal and opposite effects and hence cancel the charge.

Charge:

Non-polar molecules has overall no charge.

Shapes of Non polar bond:

Symmetrical polyatomic molecules are non-polar molecules having shape of linear, trigonal, Planar and tetrahedral.

Applications:

1- $O=C=O$ ⇒ due to E.N difference b/w Carbon and oxygen C=O bond in CO_2 is polar, but CO_2 is non-polar because dipole gets cancelled with each other.

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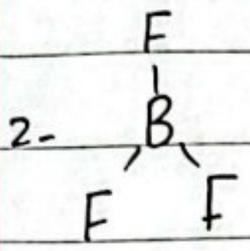
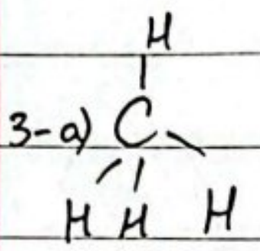
$\Rightarrow 0.7 - 0.7 = 0$

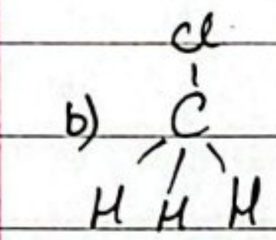
The structure is lenear and their $\mu = 0$.

2- 

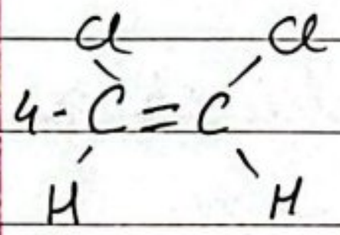
$\Rightarrow$ In trigonal planner the $\mu = 0$

3-a) 

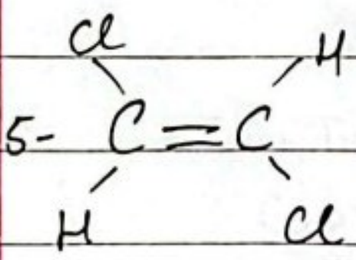
$\Rightarrow$ It is a tetrahedral structure there $\mu = 0$ because the atoms are same.

b) 

$\Rightarrow$ It is also a tetrahedral structure there $\mu \neq 0$ because the atoms are different.

4- 

$\Rightarrow$ opposite atoms are different so $\mu \neq 0$

5- 

$\Rightarrow$ opposite atoms are same so $\mu = 0$.

Formula for percantage Ionic character:

$$\% \text{ ionic character} = \frac{\mu_{obs}}{\mu_{ionic}} \times 100$$

-> below 50% = covalent

-> above 50% = ionic

* CsF have the highest ionic character.

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Concept Assessment Exercise: 3.1 (Pg 43):-

Q:1 Are the following molecules polar or non-polar? In each case give a reason for your answers.

(Electronegativity values : F = 4.0, Cl = 3.0, Br = 2.8
S = 2.5, C = 2.5, H = 2.1)

a. Chlorine, Cl_2:-

Cl — Cl $\Rightarrow$ 3.0 - 3.0 = 0

3.0 3.0

Due to electronegative difference between Cl-Cl bond is non-polar because dipole get cancelled with each other and it is linear in shape.

b. Hydrogen fluoride, HF:-

$H^{\delta+} - F^{\delta-}$ $\Rightarrow$ 4.0 - 2.1 = 1.9

2.1 4.0

It is a polar molecules as Fluorine atom is more E.N and will attract shared electron to itself and electronegative difference is not equal 0 ($\mu \neq 0$)

c. The V-Shaped molecule, Sulfur dichloride, SCl_2:-

Cl—S—Cl $\Rightarrow$ 3.0 - 2.5 = 0.5

It is a polar molecule as chlorine atom is more E.N and will attract shared electron and electronegative difference is not equal 0 ($\mu \neq 0$).

d. The tetrahedral molecule, Chloromethane, CH_3Cl.

It is a polar molecule as the difference of E.N is more and electrons are attracted.

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BOND ENERGY

Definition:

"Energy which is required to break the bond of the same type in one mole of substance is known as bond energy."

$H-H \rightarrow H+H \quad \Delta H = 436\ KJ/mol$ (Endothermic)

$H+H \rightarrow H-H \quad \Delta H = -436\ KJ/mol$ (Exothermic).

Factors:

1- **Electronegativity difference between two bonded atoms:**

The greater the E.N difference b/w two bonded atoms, the more polar will be the bond and greater will be bond energy.

Bond	Bond Energy	E.N difference b/w bonded atoms.
H-F	562	4 - 2.1 = 1.9
H-Cl	431	3 - 2.1 = 0.9
H-Br	366	2.8 - 2.1 = 0.7
H-I	299	2.5 - 2.1 = 0.4

H-F > H-Cl > H-Br > H-I

Bond energy $\propto \Delta$ electronegativity difference.

* HF has the highest value of bond energy and HI has the lowest bond energy.

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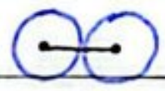
2. **Size of atom:**

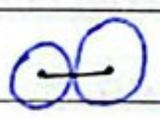
Bond energy has inverse relation with size of atom. More will be the bond energy. less will be atomic size and vice versa. This is because as the force of attraction between molecule is small so the bond energy is decrease and they can easily break.

$B.E \propto \frac{1}{Atomic\ size}$.

For e.g:



H-H



H-Br

3- **Bond Order:** (B.O)

Bond energy has direct relation with bond order. More will be the bond energy, more will be the order. As the number of bond increases in a molecule like:

C–C	B.O = 1	1.54A°	$B.E \propto B.O$
C=C	B.O = 2	1.34A°	
C≡C	B.O = 3	1.20A°	

So, C≡C > C=C > C–C

$1A° = 10^{-10} m$

4- **Bond length:**

"It is the distance between the nuclei of two covalently bonded atoms."

dependence:

It depends upon the size of atom. Smaller the size of atoms, the greater the bond energy due to shorter bond length.

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Measurment:

It is measured in A° or picometer (Pm).

$1 pm = 10^{-12} m$

Relation:

Bond energy has inverse relation with bond length.

Example:-

Bond energy of H-H = 436 $KJ mol^{-1}$

Bond energy of H-Br = 366 $KJ mol^{-1}$

* As the bond of H-H is shorter as compared to H-Br. so, size of Br is greater than hydrogen.

Bond	Bond energy ($KJmol^{-1}$)	Bond length (pm)
F-F	158	149
Cl-Cl	242	199
Br-Br	193	228
I-I	151	266

The atomic radius increases from fluorine to iodine but bond energy decreases from chlorine to iodine molecule.

V.V. Imp.

Q: Why bond energy of fluorine is less than chlorine although the bond energy decreases down the group?

Ans: Bond energy decreases from chlorine to iodine molecule, but fluorine molecule has lower value of bond energy than chlorine. It is because of the smaller size of fluorine atoms that their lone pairs repel each other, hence less amount of energy is required to break the bond between fluorine atoms.

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:F-F: → due to repulsion of e⁻s in fluorine

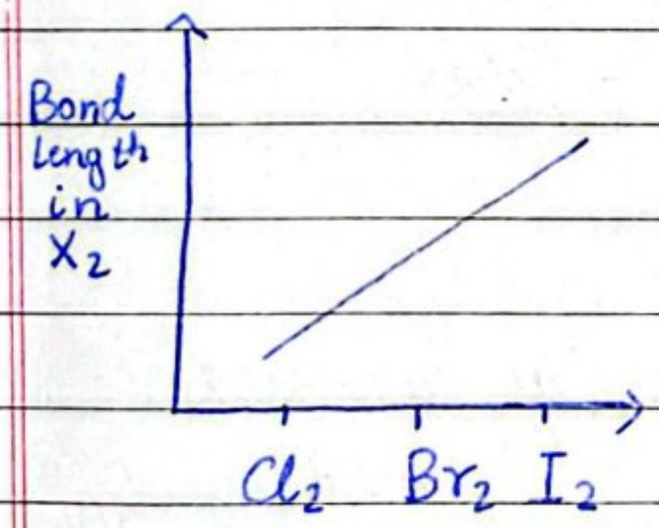
Cl-Cl > Br-Br > I-I

* Bond energy is also used to predict the reactivity of covalent molecules. Greater the bond energy stronger will be the bond and that molecule will be stable and less reactive.

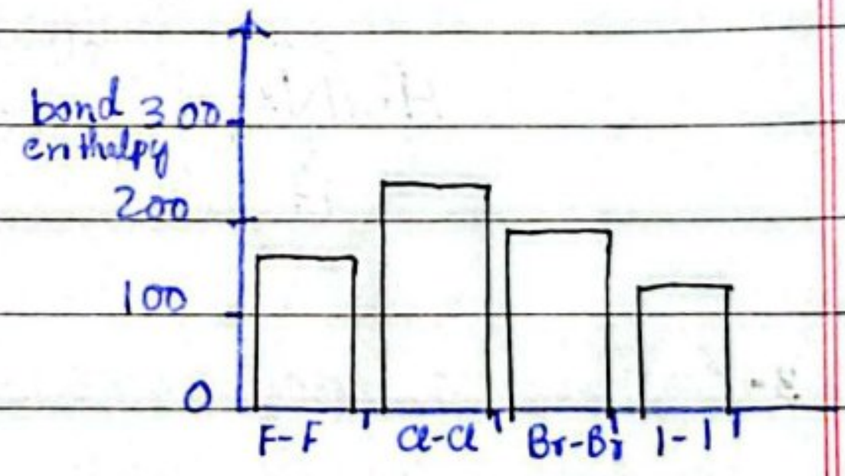
Examples:

→ In alkyl halides, alkyl fluorides are less reactive than alkyl chloride because bond energy of C-F is greater tha C-Cl.

Trend of bond length and bond energy of halogen.



a) Bond length of halogen from Cl_2 to I_2



b) Bond enthalpy of halogen from F_2 to I_2

VALENCE SHELL ELECTRON PAIR REPULSION THEORY

→ This theory is used to determine the shape and bond angle in the molecule.

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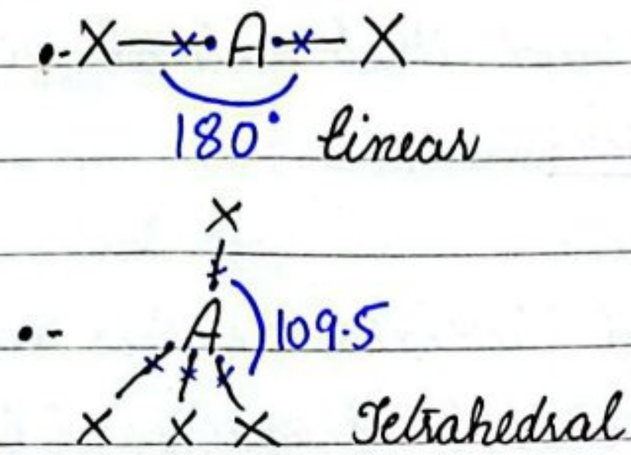
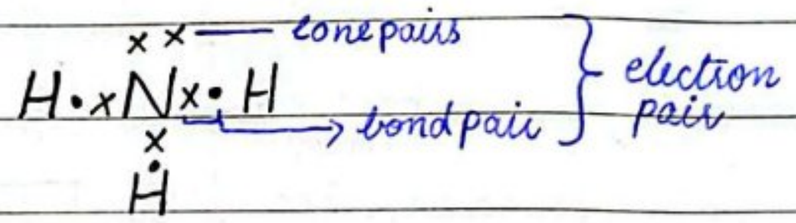
POSTULATES:

1- Central atom arrange the electron pair in such a way that there is maximum distance between the electron pair because repulsion will be minimum.



2- Repulsion order between electron pair are

lone pair-lone pair > lone pair - bond pair > bond pair - bond pair

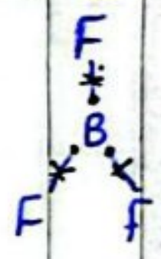
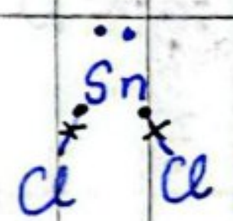
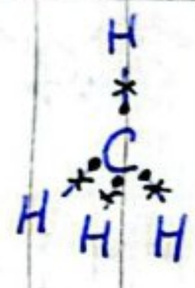


3- Lone pair occupay more space as compared to bond pair because lone pair is attracted by only nucleus and bond pair is attracted between two nuclei.

4- Double or triple bond is consider to the 1 electron pars.

5- Geometry of molecule depend upon numbers of electron pair and nature of electron pair.

SHAPES OF MOLECULES

Type	Total No's of e⁻s	bond pairs	lone pairs	Arrangment of e⁻ pair	Molecular geometry	Angle	Examples
AX_2	2	2	0	linear	linear	180°	Cl–Be–Cl, $HgCl_2$, H–C≡N, CO_2
AX_3	3	3	0	Trigonal plannar or Triangular	Trigonal Plannar	120°	 $AlCl_3$, SO_3, NO_3^{-1}
AX_2E	3	2	1	Triangular or Trigonal plannar	Bent, V-Shape, angular	<120°	 SO_2
AX_4	4	4	0	Tetrahydral	Tetrahydral	109·5	 SiH_4, CCl_4, SO_4^{-2}

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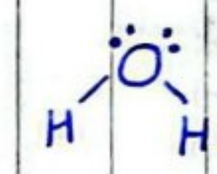
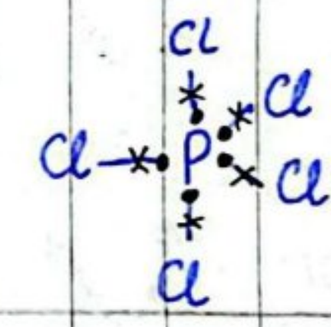
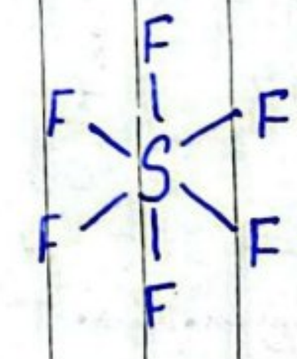
AX_3E	4	3	1	Tetrahydral	Trigonal Pyramid	<109.5 or 107.5		$\ddot{P}H_3$ PCl_3 NF_3
AX_2E_2	4	2	2	Tetrahydral	Bent, V-Shape, angular	<109.5 or 104.5		$H_2\ddot{S}$ $H_2\ddot{S}e$
AX_5	5	5	0	Trigonal bipyramidal	Trigonal bipyramidal	90, 120, 180		
AX_6	6	6	0	Tetragonal bipyramidal or octahedral	octahedral	90, 180		

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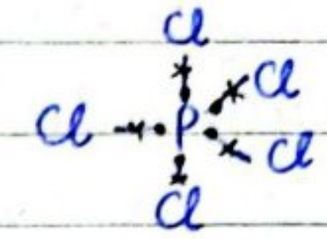
APPLICATIONS:

* Explain following structures on basis of VSEPR Theory.

1- PCl_5:



Type: AX_5

Total NO's of e⁻s: 5

Bond pairs: 5

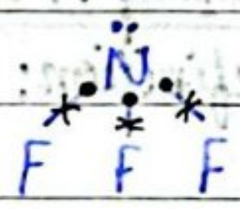
Lone pairs: 0

Arrangment of e⁻s: Trigonal bipyramidal

Molecular geometry: Trigonal bipyramidal

Angle: 90, 120, 180

2- NF_3:



Type: AX_3E

Total NO's of e⁻s: 4

Bond pairs: 3

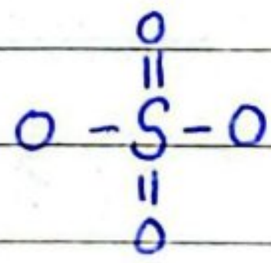
Lone pairs: 1

Arrangment of e⁻s: Tetrahydral

Molecular geometry: Trigonal pyramid

Angle: < 109.5 or 107.5

3- SO_4^{-2}



Type: AX_4

Total NO's of e⁻s: 4

Bond pairs: 4

Lone pairs: 0

Arrangment of e⁻s: Tetrahydral

Molecular geometry: Tetrahydral

Angle: 109.5

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VALENCE BOND THEORY

Proposed by:

This theory was proposed by Heitler and London (1927) and later developed by pauling. It has successfully explained bond energies, bond lengths and shapes of the covalent molecules.

Explanation:

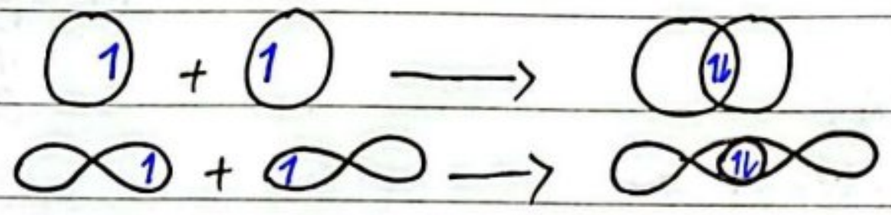
In valence Bond Theory, covalent bonds are formed by the overlapping of partially filled atomic orbital of one atom with a partially filled atomic orbital of the other. Their identity was retained by the two overlapping atomic orbitals.

Overlap Meaning:

"The term overlap means that the two orbitals share the common region in space."

Main Points of VBT:

1- The bond between two atoms is formed by the overlap of half-filled atomic orbitals of two atoms. These atomic orbitals retain their identity.



2- Electrons of overlapping orbitals have opposite spin.

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3- The number of bonds formed is equal to the number of unpaired electrons present in the outermost shell of the atom.

$N \equiv N \rightarrow 1s^{1v}, 2s^{1v}, 2p_x^{1}, 2p_y^{1}, 2p_z^{1}$

4- By the overlapping of two orbitals overlap; a single bond which is sigma bond is formed. By: the overlapping of additional orbitals, multiple bonds are formed (double and triple).

5- In order to form a bond, the overlapping orbitals must have the same symmetry with respect to the bond axis.

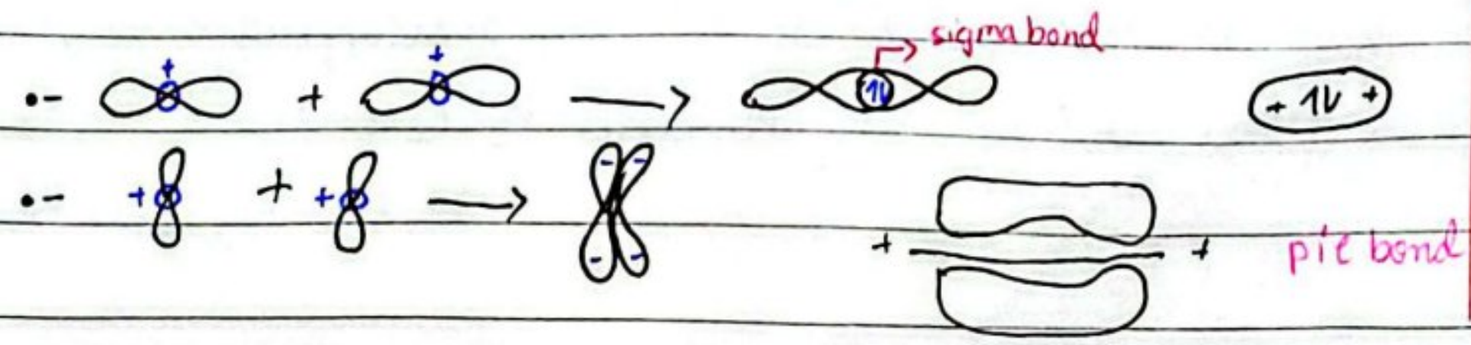
- definition of bond axis:

"A line joining the nuclei of two bonded atoms is called bond axis."

- ∞ + ∞ → ∞∞ (can overlap)
- 8 + 8 → 88 (can overlap)
- ∞ + 8 → – (cannot overlap)

6- Energy is released by the overlapping of orbitals. The greater the overlap between the orbitals, the greater is the energy released and the stronger will be the bond formed.

* Sigma bond is stronger than the pie bond.



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Types of Overlapping and Nature of Covalent Bond:

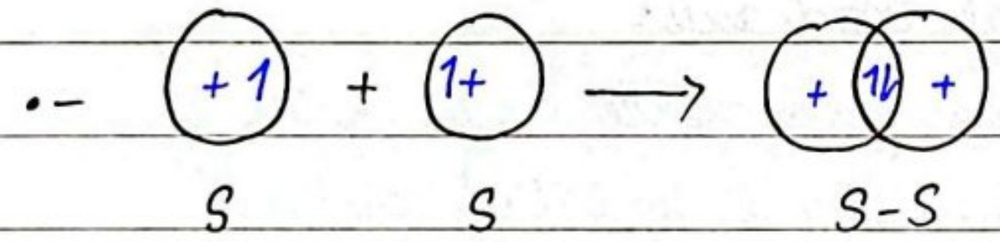
The two main types of covalent bonds obtained by overlapping.

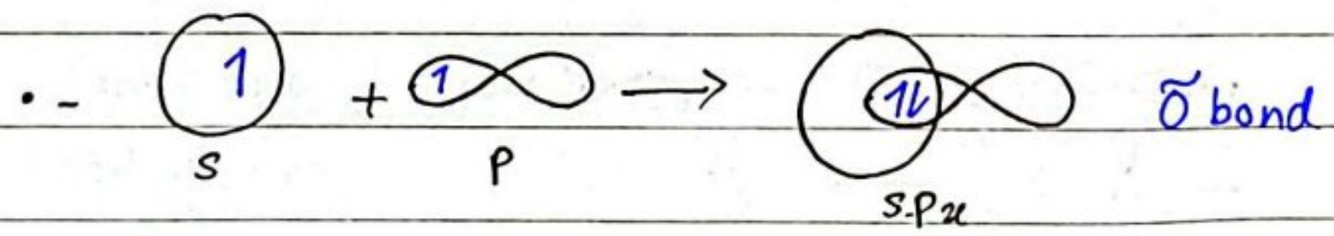
- Sigma bond (σ)
- Pi Bond (π)

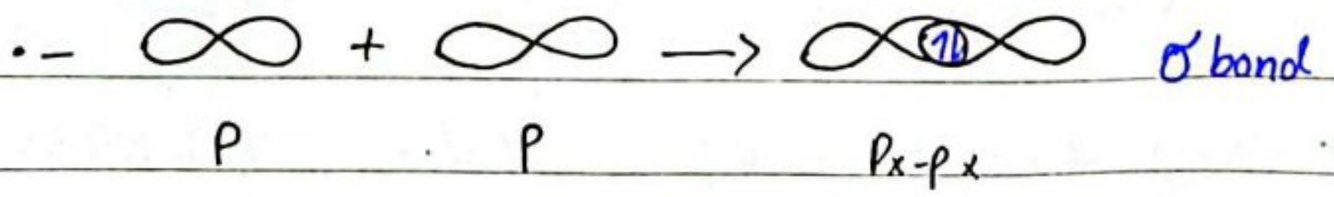
Sigma Bond:

Def: "When there is a single bond between two atoms, the bond will be sigma bond."

* It is formed by the head on overlapping of atomic orbitals.
* It is also called linear overlapping.
* It is formed between S-S, S-Px, S-Py, S-Pz, Px-Px overlap.





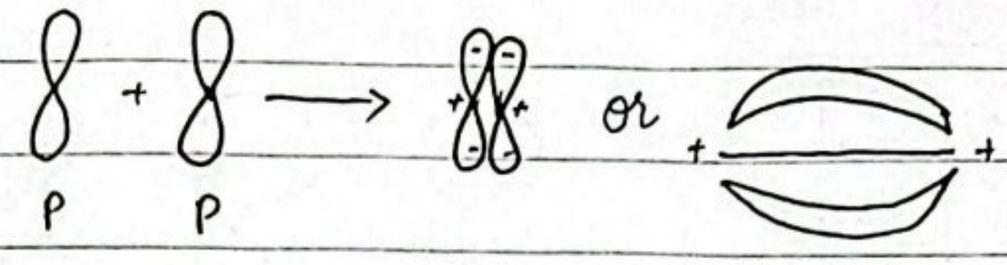


Pi Bond:

Def: "Pi bond is formed by the parallel overlapping of P orbitals."

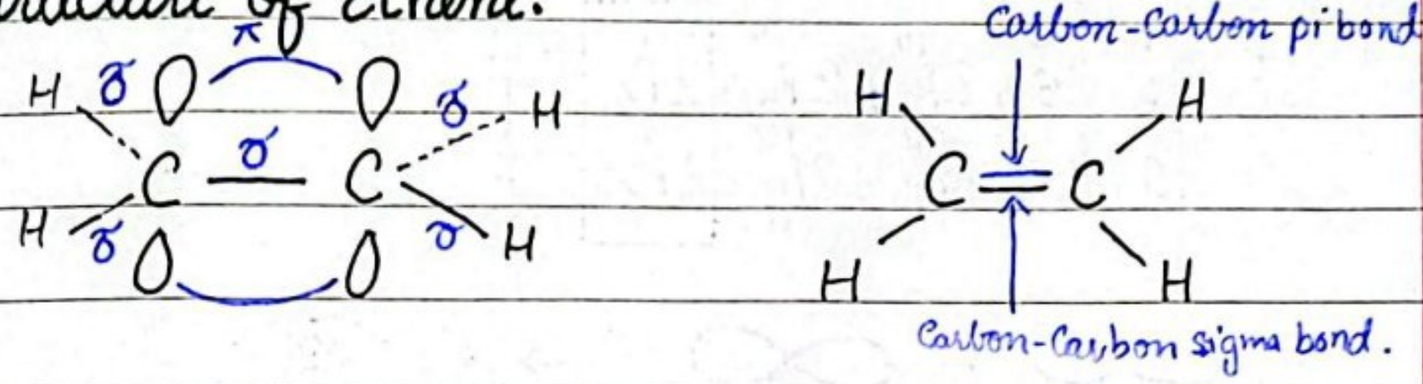
* It is also called side way overlapping.
* It is between P_y-P_y and P_z-P_z

Date: _____ Day: _____



⇒ When there is double bond between two atoms, one should be sigma and the other should be Pi bond. When there is triple bond between two atoms, there is one sigma and two Pi bond

Structure of Ethene:



Imp

Q: Why sigma bond is stronger than Pi bond?

Ans: Sigma bond is stronger than Pi bond because in sigma bond electron density is in between the line joining the two nuclei while in pi bond electron density is above and below the plane. Hence it is more diffuse bond and less amount of energy is required to break this bond.

Application of Valence Bond Theory:

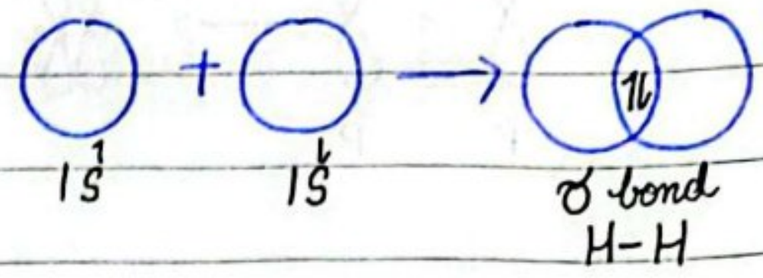
1- Formation of H_2 molecule:

Electronic configration of hydrogen atom is $1s^1$. Each half-filled s orbital of hydrogen atom overlaps to form H-H bond where electon density is in between the line joining the two nuclei so bond formed is a sigma bond (σ).

Date: _____ Day: _____

$_1H = [1s^1]$ H

$_1H = [1s^1]$ H

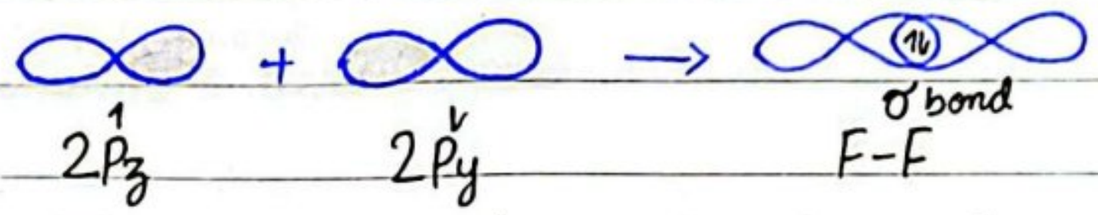


2- Formation of F_2 Molecule:

The electronic configration of fluorine is $1s^2, 2s^2, 2p^5$. Half-filled $2P_z$ orbital of each fluorine atom head on overlap to form F-F bond which is sigma bond.

$_9F = 1s^{2}, 2s^{2}, 2p_x^{2}, 2p_y^{2}, [2p_z^{1}]$ F

$_9F = 1s^{2}, 2s^{2}, 2p_x^{2}, 2p_y^{2}, [2p_z^{1}]$ F

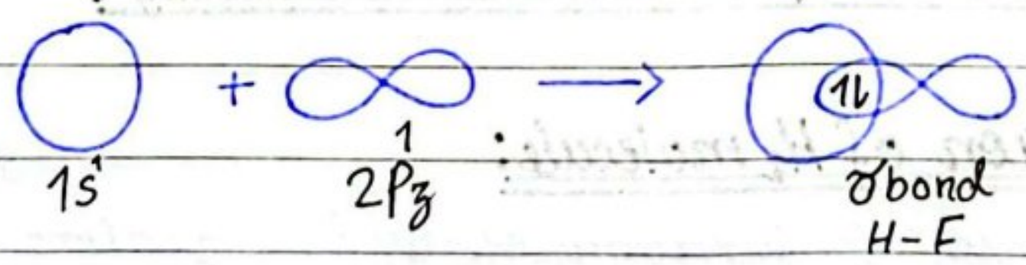


3- Formation of HF Molecule:

The electronic configration of F atom is $1s^2, 2s^2, 2p^5$. and H atom is $1s^1$. In the formation of HF molecule, the half filled 1s orbital of H atom overlaps with the half filled $2P_z$ orbital of F to form H-F σ bond.

$_1H =$ $[1s^1]$ H

$_9F = 1s^{2}, 2s^{2}, 2p_x^{2}, 2p_y^{2}, [2p_z^{1}]$ F



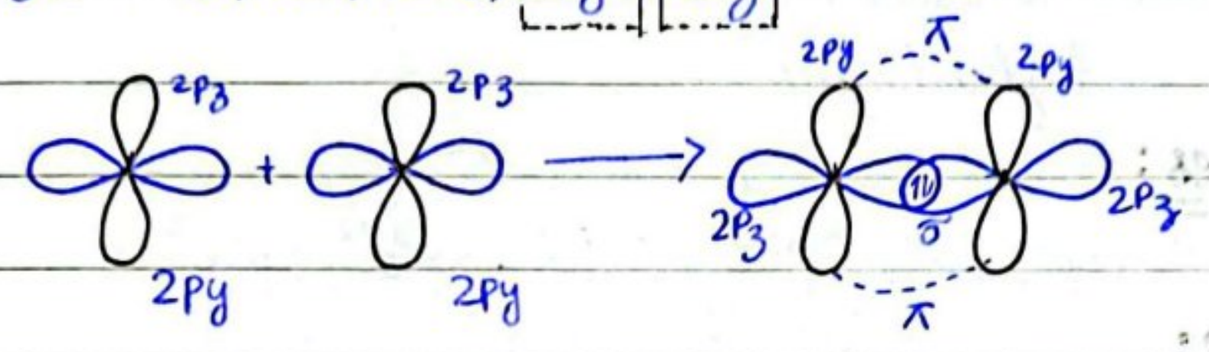
Date: _____ Day: _____

4. Formation of O_2 Molecules:

The oxygen atom contains two half-filled $2p_y$, $2p_z$ orbital so according to VBT they form two bonds. One is due to head to head overlap that is sigma bond and other due to parallel overlap of half filled porbital is pi bond.

$_8O = 1s^{\uparrow\downarrow}, 2s^{\uparrow\downarrow}, 2p_x^{\uparrow\downarrow}, 2p_y^{\uparrow}, 2p_z^{\uparrow}$ O

$_8O = 1s^{\uparrow\downarrow}, 2s^{\uparrow\downarrow}, 2p_x^{\uparrow\downarrow}, 2p_y^{\uparrow\downarrow}, 2p_z^{\uparrow}$ O

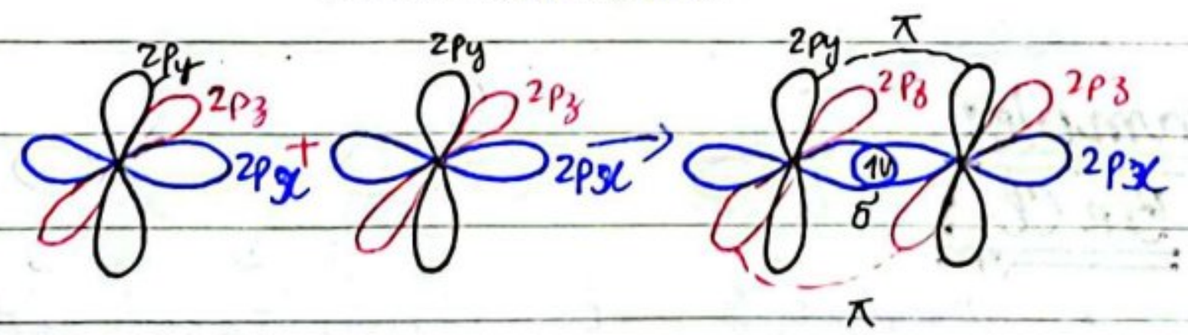


5. Formation of N_2 Molecules:

The nitrogen atom contains three half-filled p orbitals. Half-filled $2p_x$ orbitals of nitrogen atoms head on overlap with each other and sigma bond is formed, half-filled orbital of $2p_y$ and $2p_z$ of nitrogen atom are parallel so two pi-bonds are formed.

$_7N = 1s^{\uparrow\downarrow}, 2s^{\uparrow\downarrow}, 2p_x^{\uparrow}, 2p_y^{\uparrow}, 2p_z^{\uparrow}$ N

$_7N = 1s^{\uparrow\downarrow}, 2s^{\uparrow\downarrow}, 2p_x^{\uparrow}, 2p_y^{\uparrow}, 2p_z^{\uparrow}$ N



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HYBRIDIZATION or ORBITAL HYBRIDIZATION

Definition:

Process in which orbital of different energy or different shape are mix into eachother and form a new orbital that is of equal energy and same shape is called hybridization."

Types:

$sp, sp^2, sp^3, sp^3d, sp^2d, sp^3d^2, sp^3d^3$

Rule:

No of sigma bonds + lone pairs = 2 = sp = AX_2

= 3 = sp^2 = AX_3, AX_2E

= 4 = sp^3 = AX_4, AX_3E, AX_2E_2

SP-Hybridization:

"Mixing of one 's' and one 'p' i.e. P_x forming two new orbitals (hybrid or orbitals) of equal energy and same shape is called sp-hybridization."

→* It contains 50% of s-character and 50% p-character.

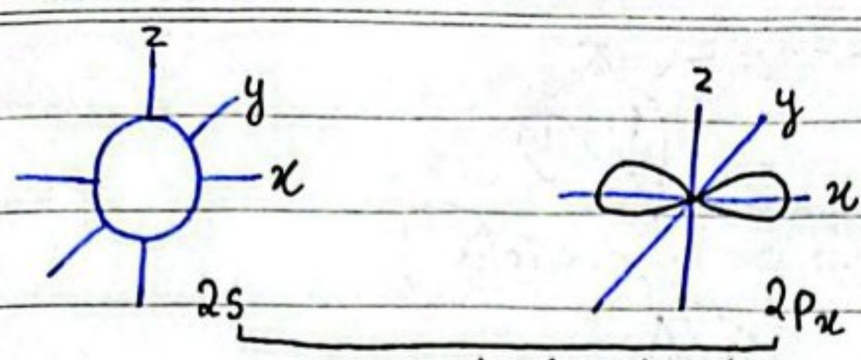
Example:

1- **$BeCl_2$:**

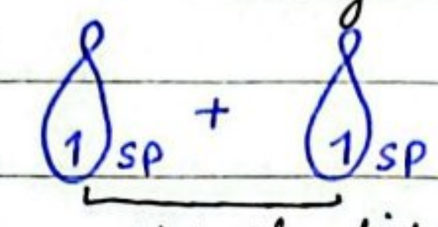
- Ground state $_4Be = \underline{1s^{\uparrow\downarrow}}, \underline{2s^{\uparrow\downarrow}, 2p_x, 2p_y, 2p_z}$ (K, L)
- Excited State $_4Be = 1s^{\uparrow\downarrow}, 2s^{\uparrow}, 2p_x^{\uparrow}, 2p_y^{\cdot}, 2p_z^{\cdot}$
- Hybrid state $_4Be = 1s^{\uparrow\downarrow}$ [1] [1]

Sp Sp

Date: _____ Day: _____



unhybrid orbitals



Hybrid orbitals

$_{11}Cl = 1s^2, 2s^2, 2p^6, 3s^2, 3p_x^{1v}, 3p_y^{1v}, 3p_z^1$

Molecular geometry = Linear

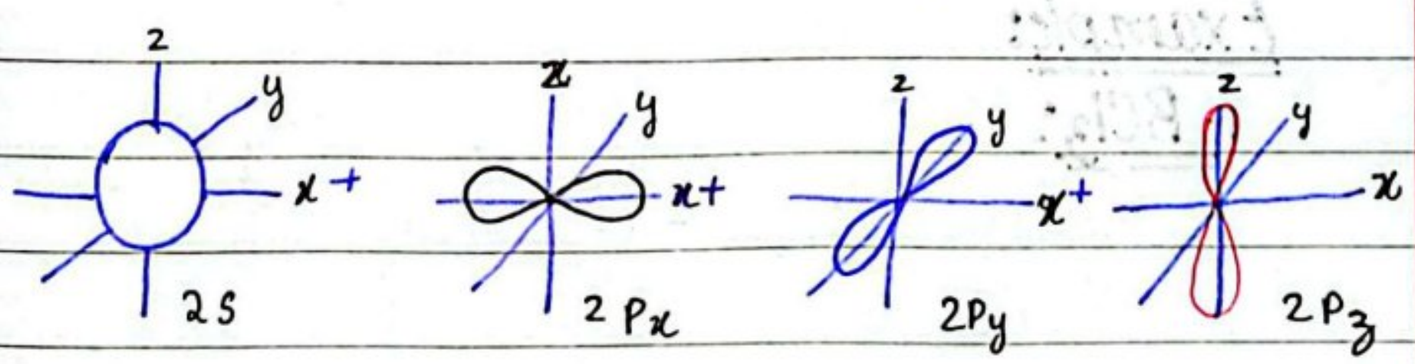
Angle = 180°

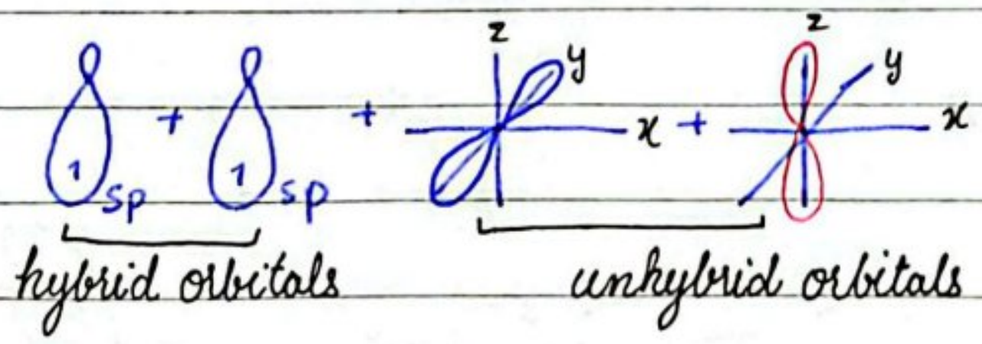
Sigma bond = 2

2- H−C≡C−H (Ethyne or Acetylene)

- Ground state $_6C = \underbrace{1s^{1v}, 2s^{1v}}_{K}, \underbrace{2p_x^1, 2p_y^1, 2p_z^1}_{L}$
- Excited state $_6C = 1s^{1v}, 2s^1, 2p_x^1, 2p_y^1, 2p_z^1$
- Hybrid state $_6C = 1s^{1v}, \boxed{1}\boxed{1}, 2p_y^1, 2p_z^1$

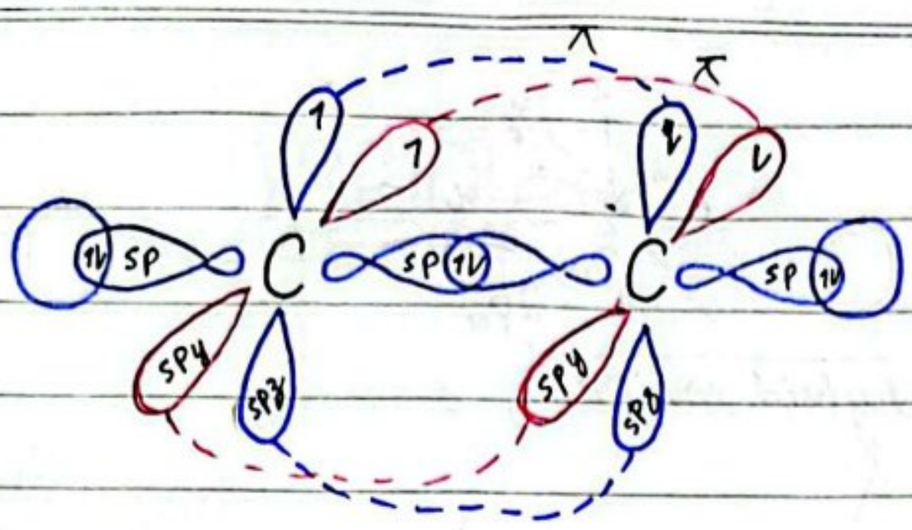
sp sp





hybrid orbitals unhybrid orbitals

Date: _____ Day: _____



Molecular geometry = Linear

Angle = 180°

Sigma bond = 3

Pi bond = 2

SP^2 Hybridization:

"Mixing of one 's' and two 'p' i.e. Px & Py forming three new orbitals (hybrid orbitals) of equal energy energy and same shape is called sp^2-hybridization."

→ * It contains 33.3% s-character and 66.7% p-character.

Example:

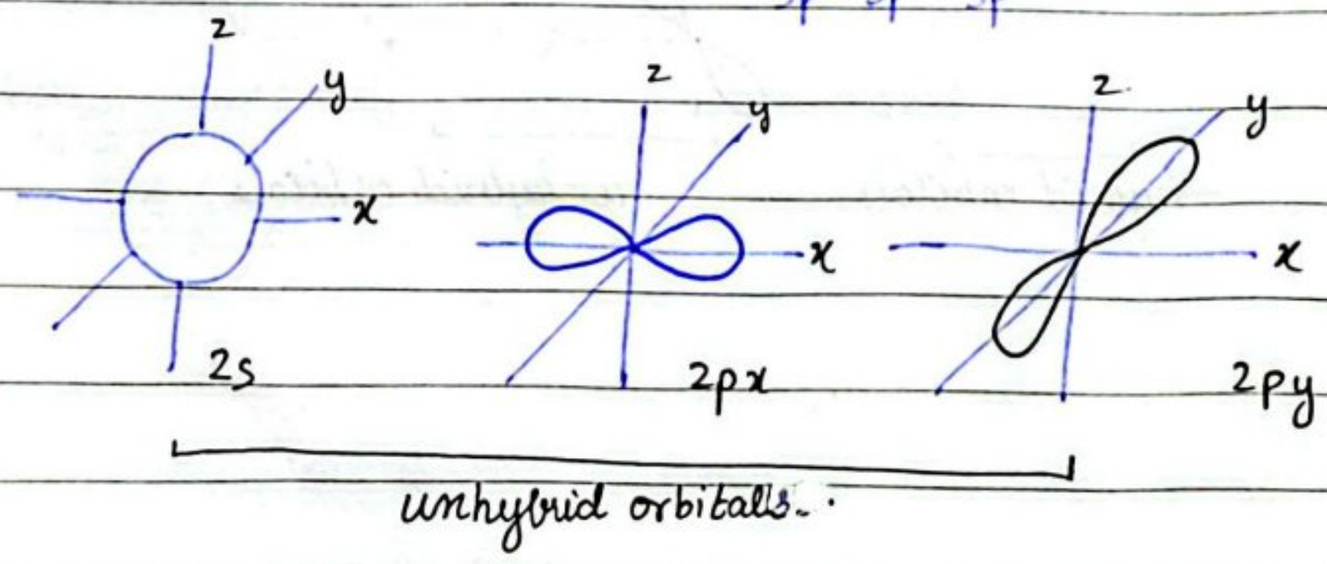
1- BCl_3:

Ground state: $_5B = 1s^2, 2s^2, 2p_x^1, 2p_y^0, 2p_z^0$

Excited state $_5B = 1s^2, 2s^1, 2p_x^1, 2p_y^1, 2p_z^0$

Hybrid state $_5B = 1s^2,$ [1] [1] [1]

sp^2 sp^2 sp^2



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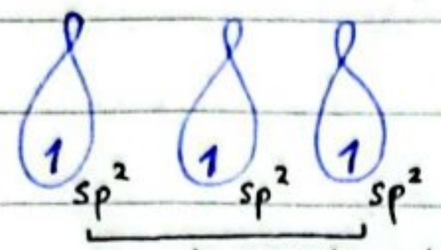
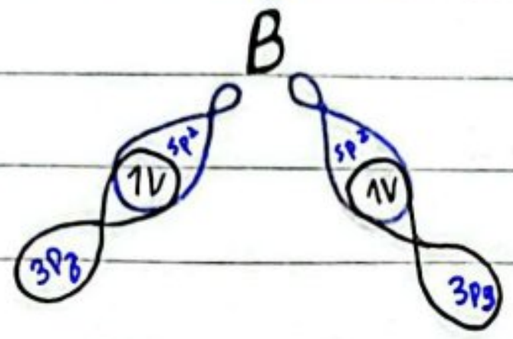
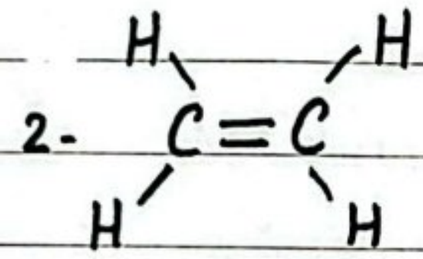


hybrid orbitals

$_{17}Cl = 1s^2, 2s^2, 2p^6, 3s^2, 3p_x^{1v}, 3p_y^{1v}, 3p_z^1$



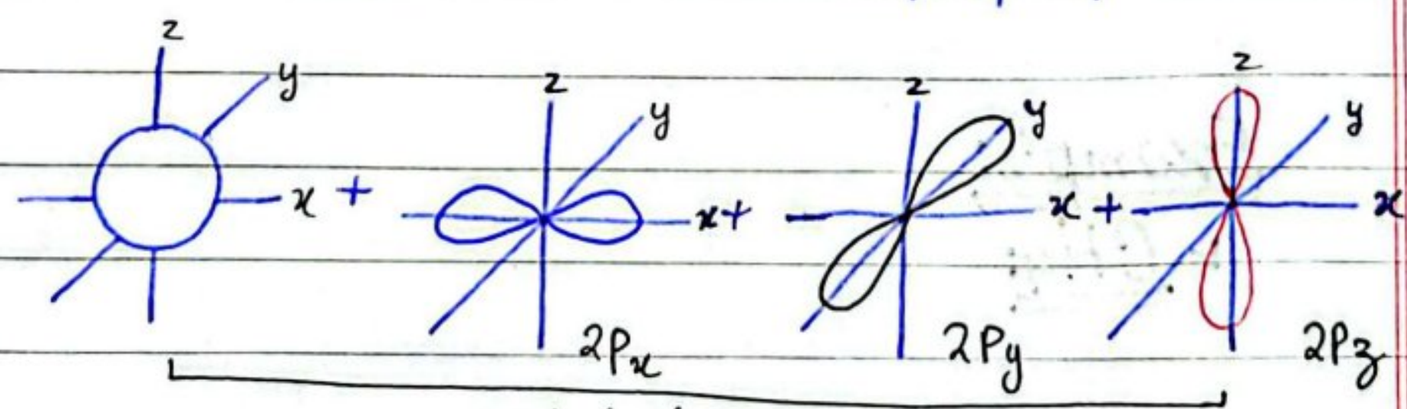
Molecular geometry = Trigonal Planar
Angle = 120°
sigma = 3

2- 

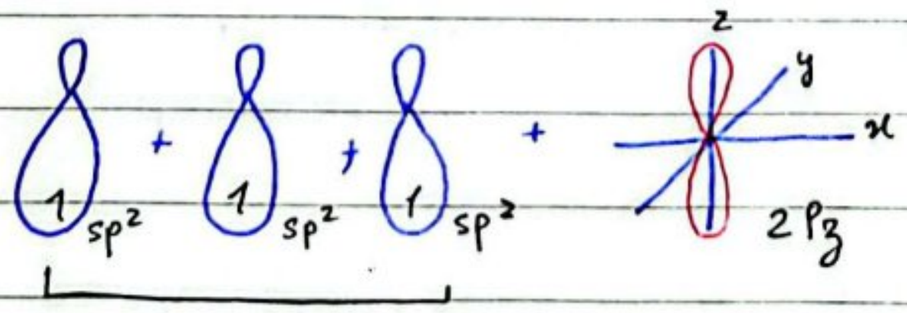
(Ethene / Ethylene)

- Ground state. $_6C = 1s^{1v}, 2s^{1v}, 2p_x^1, 2p_y^1, 2p_z^0$
- Excited state $_6C = 1s^{1v}, 2s^1, 2p_x^1, 2p_y^1, 2p_z^1$
- Hybrid state $_6C = 1s^{1v}, [1|1|1], 2p_z^1$

$sp^2\ sp^2\ sp^2$



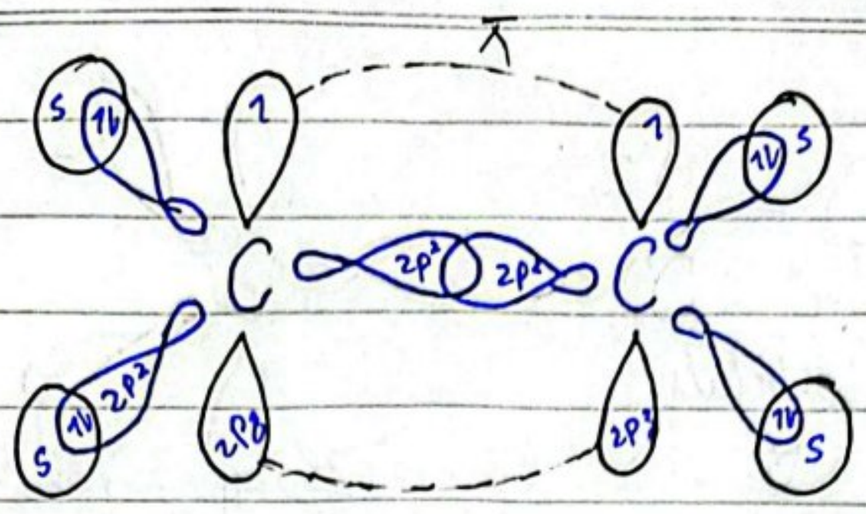
unhybrid orbitals



hybrid orbitals

Page:28

Date: Day:



Molecular geometry = Trigonal Planner

Angle = 120°

Sigma bond = 5

Pi- bond = 1

SP^3 - Hybridization:

"Mixing of one s and 3p that is P_x, P_y, P_z forming 4 new orbital of equal energy and same shape is called sp^3 hybridization."

→* It contains 25% s-character and 75% p-character.

Example:

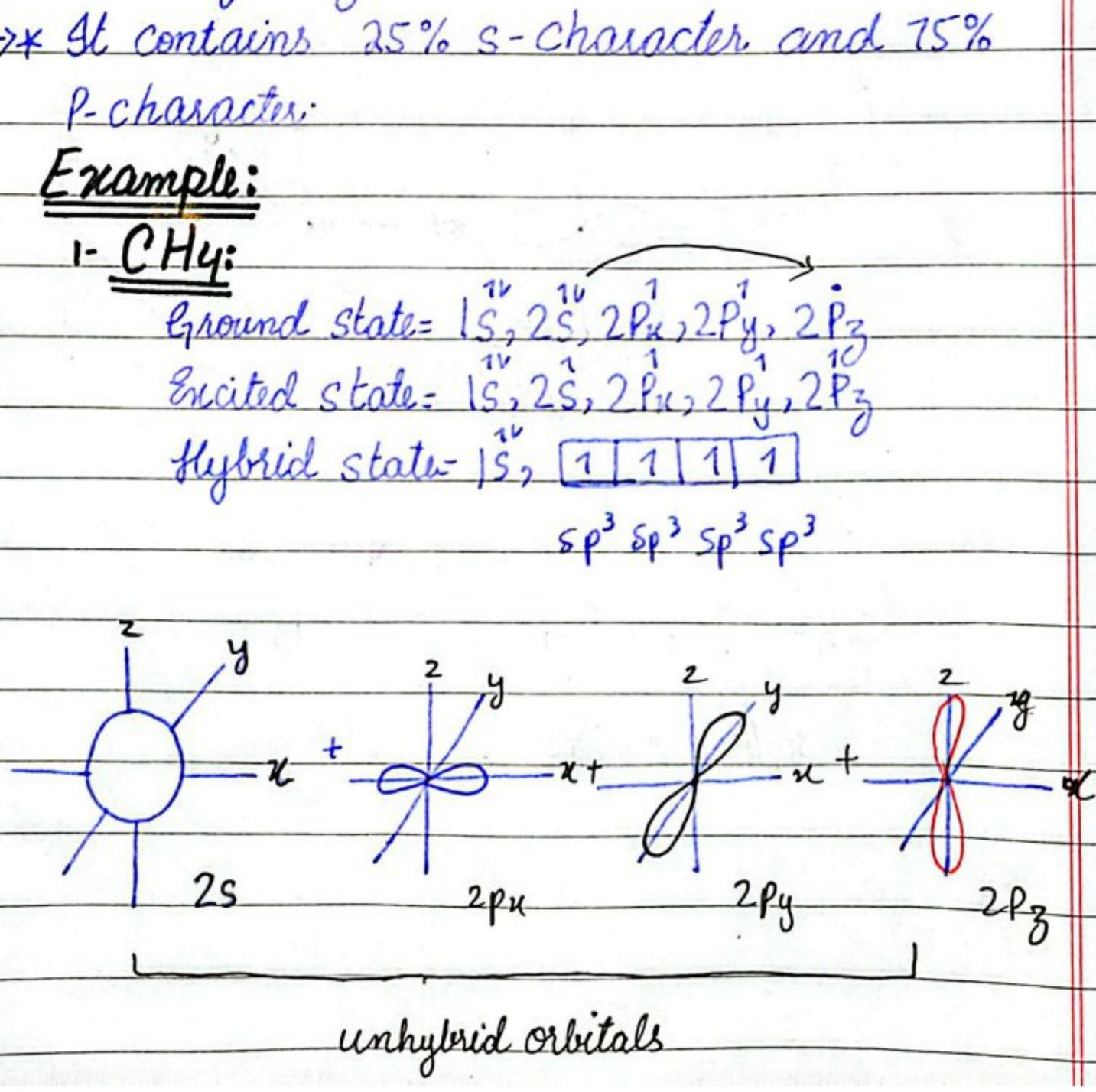
1- CH_4:

Ground state = $1S^{2}, 2S^{2}, 2P_x^{1}, 2P_y^{1}, 2P_z$

Excited state = $1S^{2}, 2S^{1}, 2P_x^{1}, 2P_y^{1}, 2P_z^{1}$

Hybrid state = $1S^{2}$, [1 | 1 | 1 | 1]

$sp^3\ sp^3\ sp^3\ sp^3$



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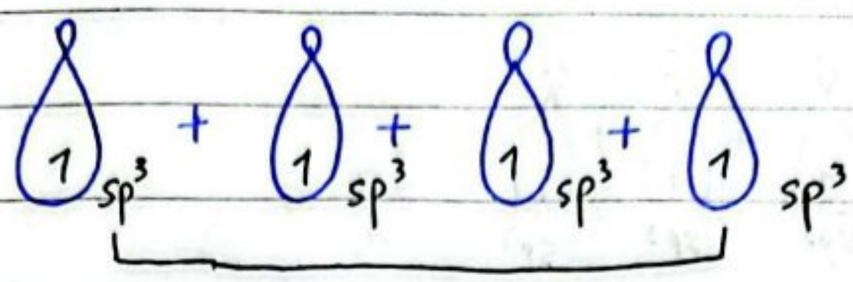
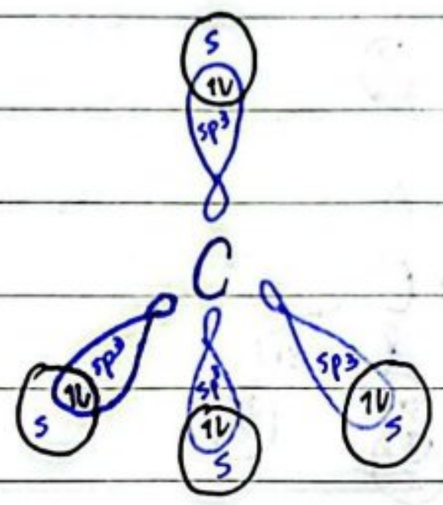


hybrid orbitals



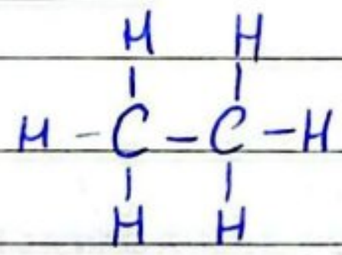
$_1H = 1s^1$

Molecular geometry = Tetrahydral

Angle = 109.5°

Sigma bond = 4

2- C_2H_6 (Ethane):

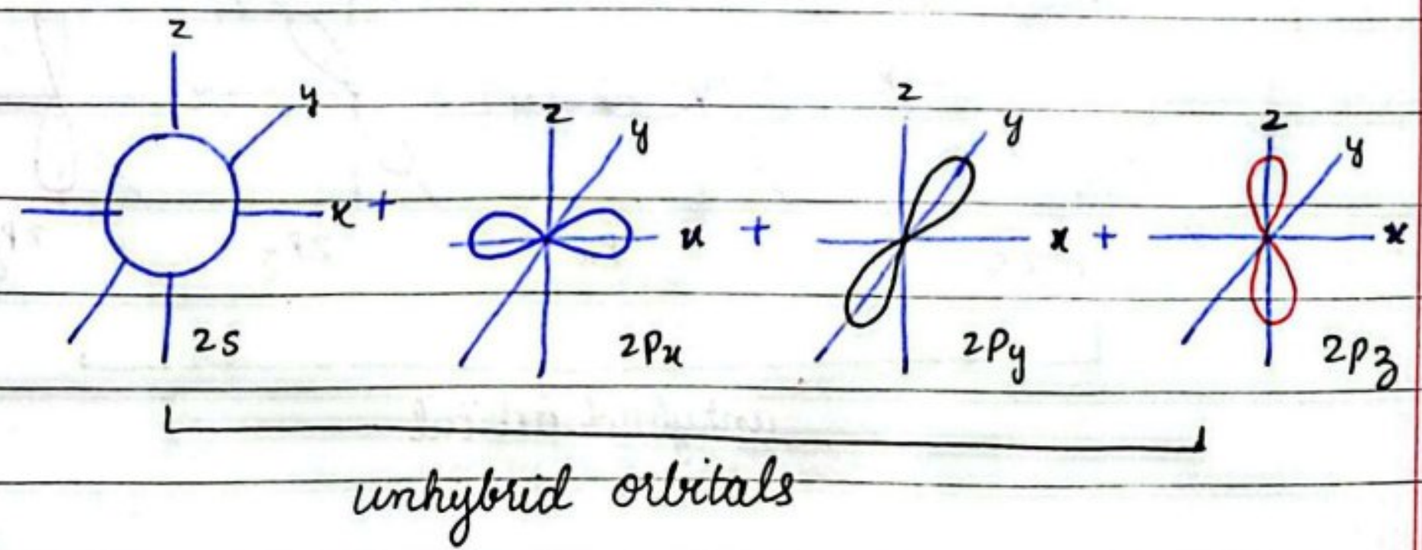


Ground state = $_6C = 1s^{1V}, 2s^{1V}, 2P_x^1, 2P_y^1, 2P_z^1$

Excited state = $_6C = 1s^{1V}, 2s^1, 2P_x^1, 2P_y^1, 2P_z^1$

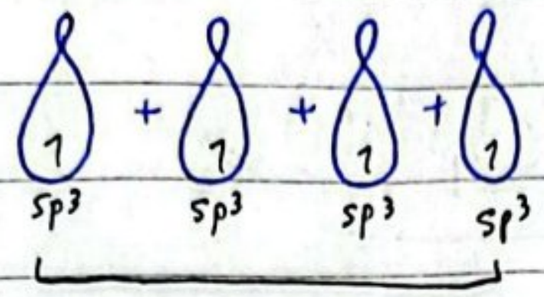
Hybrid state = $_6C = 1s^{1V}$, [1][1][1][1]

sp^3 sp^3 sp^3 sp^3

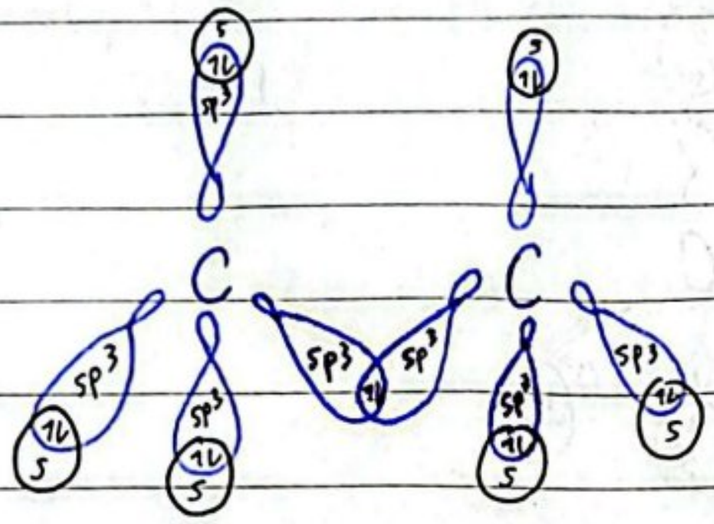


unhybrid orbitals

Date: Day:



hybrid orbitals

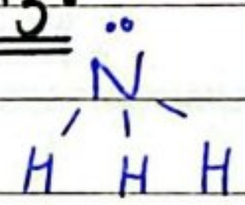


Molecular geometry = Tetrahydral

Angle = 109.5

Sigma bond = 7

3- NH_3 :

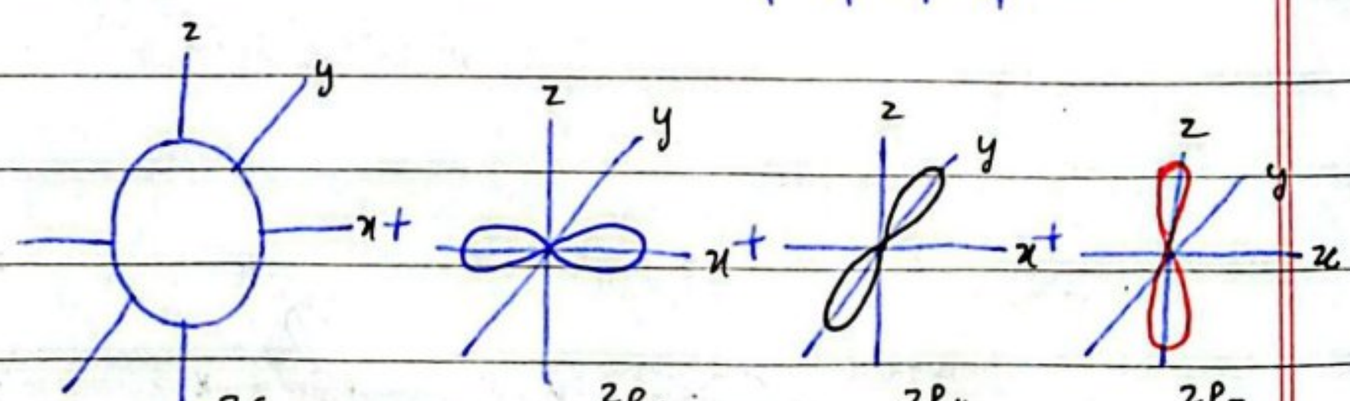


Ground state/ $= {}_7N = 1s^2, 2s^2, 2p_x^1, 2p_y^1, 2p_z^1$

Excited State

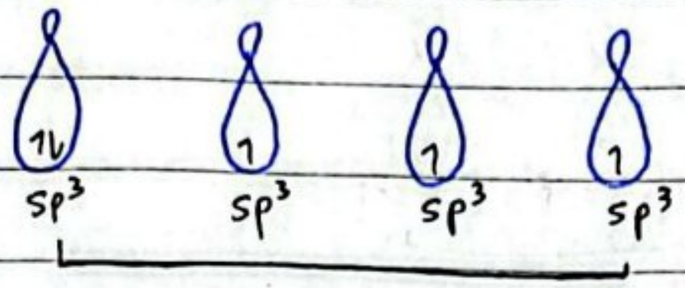
Hybrid state $= {}_7N = 1s^1$, [1v | 1 | 1 | 1]

$sp^3\ sp^3\ sp^3\ sp^3$

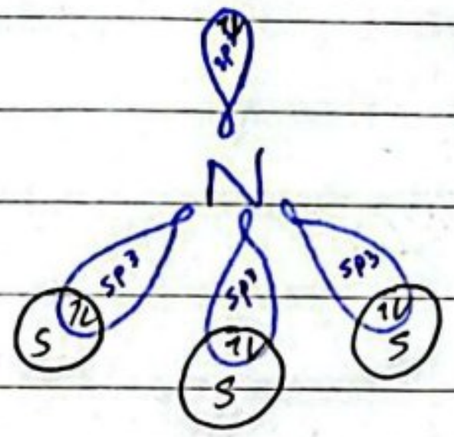


unhybrid orbital

Date: _____ Day: _____



hybrid orbitals



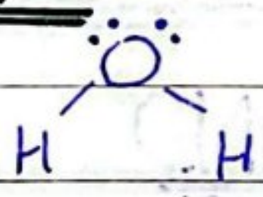
Molecular Geometry = Trigonal Pyramidal

Angle = 107°

Sigma bond = 3

lone pair = 1

4- H_2O:

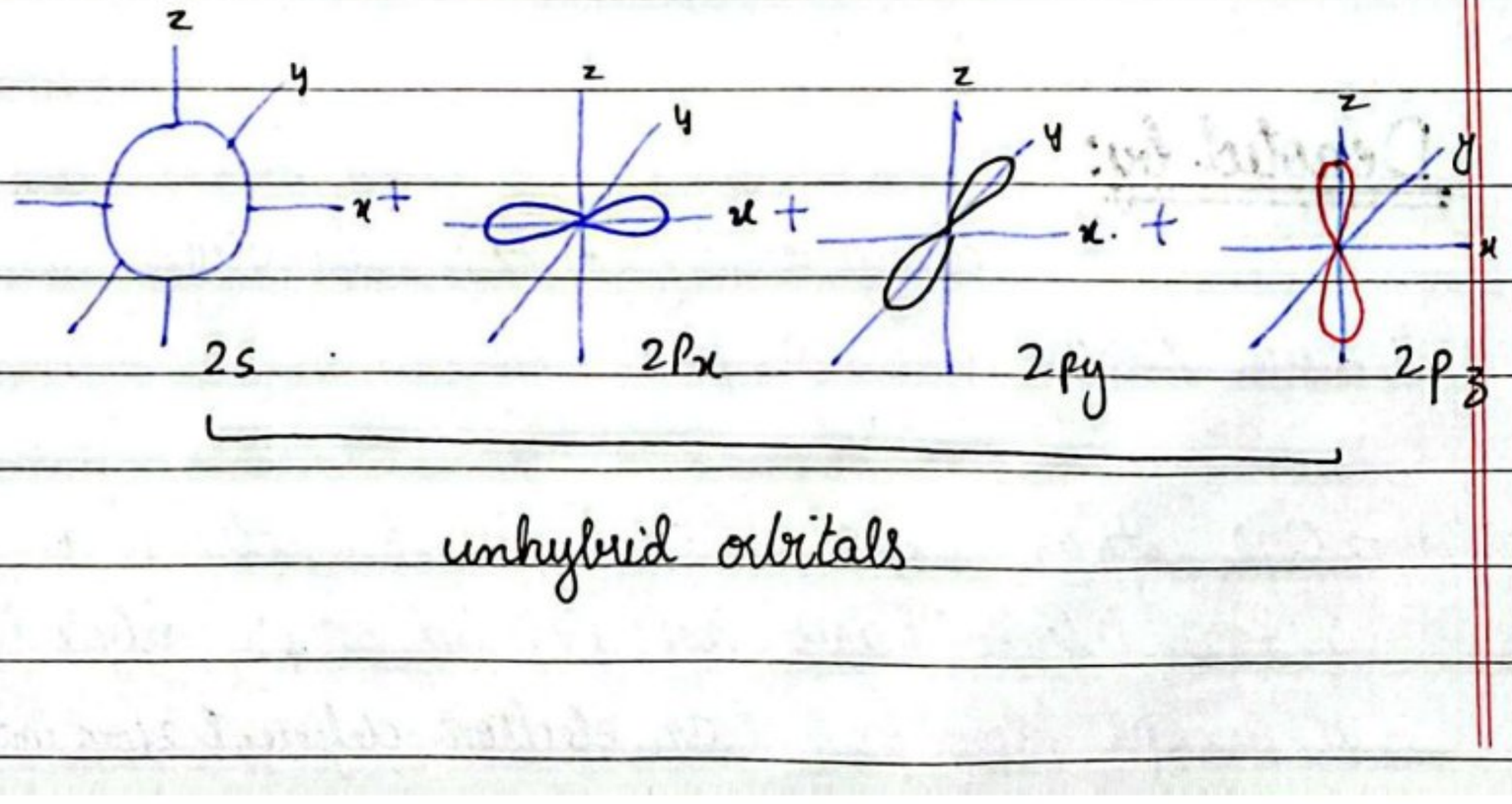


Ground state / = $_8O$ = $1S^{2}, 2S^{2}, 2P_x^{2}, 2P_y^{1}, 2P_z^{1}$

Excited state

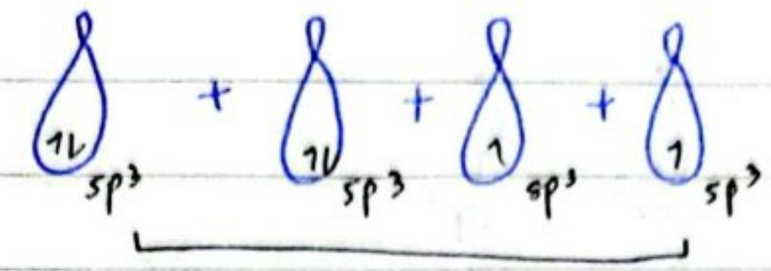
Hybrid state = $_8O$ = $1S^{2}$, | 2 | 2 | 1 | 1 |

sp^3 sp^3 sp^3 sp^3

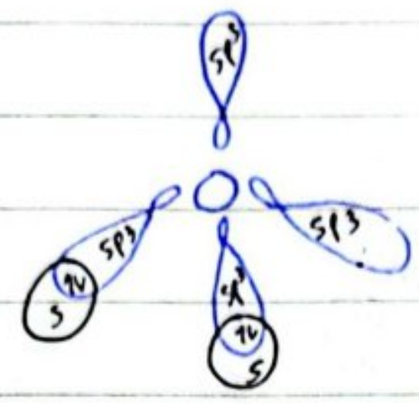


unhybrid orbitals

Date: Day:



hybrid orbitals



Molecular geometry = V-Shape

Angle = 104.2

Sigma bond = 2

lone pairs = 2

CO-ORDINATE / DATIVE

COVALENT BOND

Definition:

"A co-ordinate covalent bond is formed by sharing of electrons or shared pair of electrons and denoted by one of bond atom."

Denoted by:

It is denoted by (⟶). diriction is from donor to acceptor.

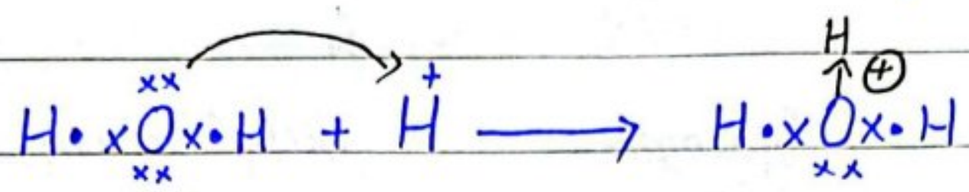
*→ One atom have a lone pair of electron and second atom have an unfilled empty orbital to accept lone pair (an electron deficient atom/molecule)

Date: _____ Day: _____

Example:

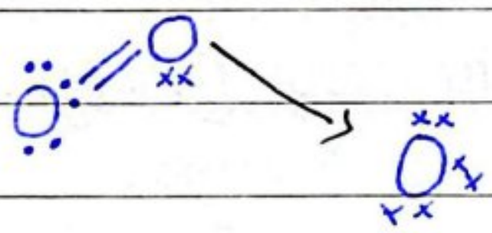
1- Hydronium ion :-

water containing lone pair of electrons donate a pair of electrons and form dative bond with the electron deficient hydrogen ion.

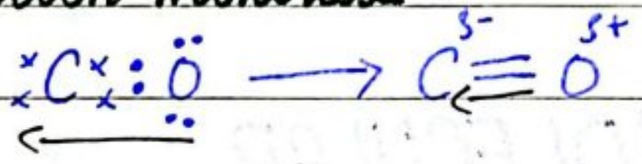


2- Ozone molecule (O_3):

central oxygen atom donates its electron pair with other oxygen atom and form dative bond.



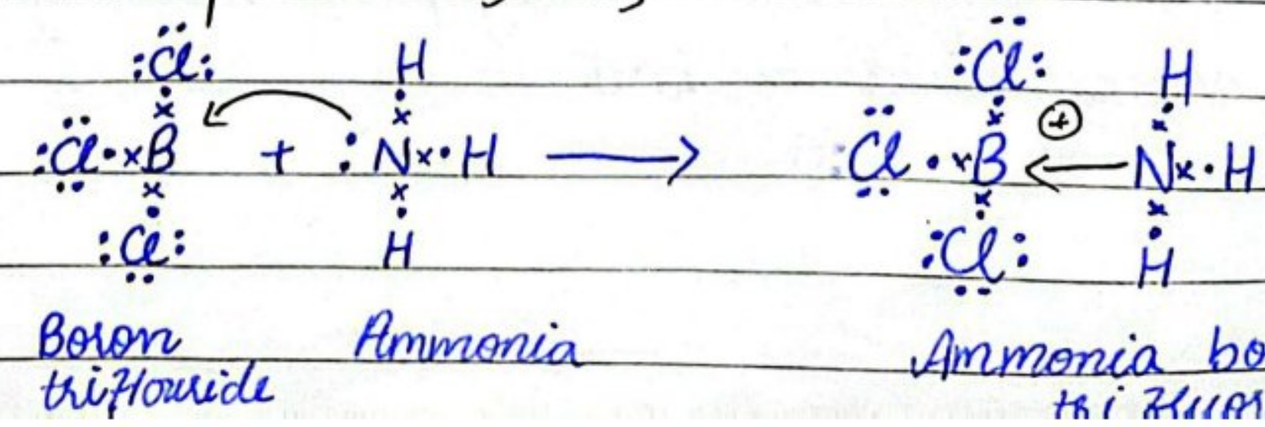
3- Carbon monoxide :



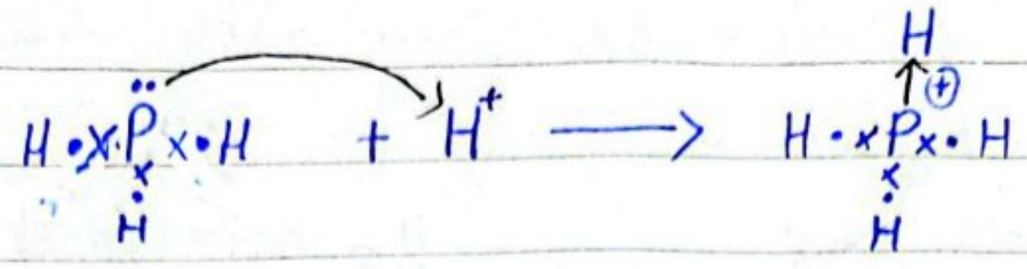
Concept Assessment Exercise 3.3 : (Page 62)

1- Draw dot and cross diagram to show the formation of co-ordinate bond between the following.

i- Boron trifluoride BCl_3 and Ammonia NH_3 to form the compound Cl_3BNH_3



ii- Phosphine PH_3 and a hydrogen ion H^+ to form the ion PH_4^+



2- Draw the displayed formulae of the products formed in part a. Show the Co-ordinate bond by an arrow.

1- Cl_3BNH_3:

$$BCl_3 + NH_3 \longrightarrow [BCl_3 \leftarrow NH_3]^+$$

2- PH_4^+

$$PH_3 + H^+ \longrightarrow [PH_3 \rightarrow H^+]^+$$

INTER MOLECULAR FORCES

Definition:

Intermolecular forces are those forces which are present between the covalent molecules.

Physical Properties:

Melting point, boiling point, density etc are explained by IMF.

Date: Day:

Types of Intermolecular Forces:-

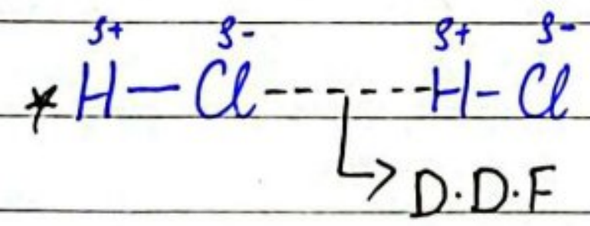
There are 3 types of IMF.

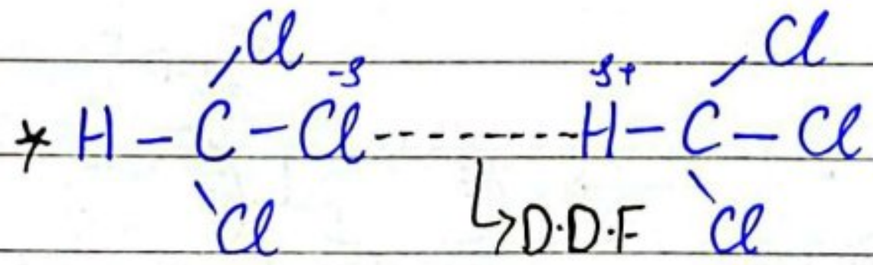
- Permanent dipole-dipole Forces (DDF)
- Instantaneous dipole induced dipole forces / London dispersion forces.
- Hydrogen Bonding

1- Permanent dipole-dipole Forces (DDF)

def: "Force of attraction between positive pole of one molecule and negative pole of other molecule is called dipole-dipole force."

- Between two polar molecules.
- 1% effective as covalent bond.
- Distance between molecule $\propto \frac{1}{DDF}$
- $\Delta E.N \propto D.D.F$
- heat of sublimation (ΔH_s), Melting point, Boiling point, heat of vaporization (ΔH_v) etc are physical properties.





Date: _____ Day: _____

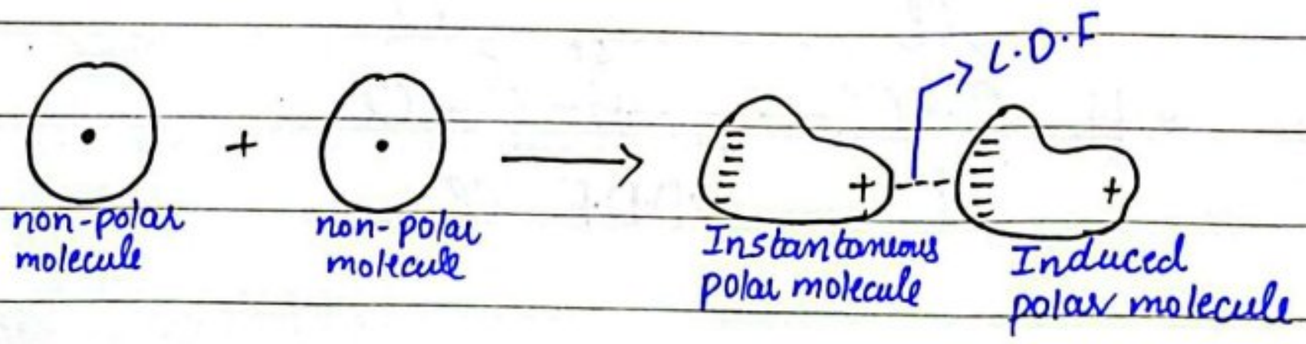
2- Instantaneous dipole - Induced dipole forces:

def:- "Force of attraction between Instantaneous polar molecule and Induced polar molecule are called L.D.F."

- Also known as London dispersion forces, Vander waal forces.
- The most weak type of Intermolecular forces.
- Two non-polar molecules.
- Even in molecules with no polar bonds, they are temporary dipoles because of uneven electron distribution due to the constant movement of electrons. This induces a temporary dipole induced dipole forces of attraction.

Factors:

1. Atomic size $\propto$ L.D.F
2. Molecular size $\propto$ L.D.F
3. Molecular mass $\propto$ L.D.F
4. No of e⁻s in molecule $\propto$ L.D.F / size of e⁻s cloud
5. Polarizability $\propto$ L.D.F



F_2, Cl_2] gases — C_1-C_4 gas

Br_2] liquid — C_5-C_{17} liq

I_2] solid — C_{18} --- solid

He, Ne, Ar, Kr, Xe, Rn] m.B & B.P increases down the group

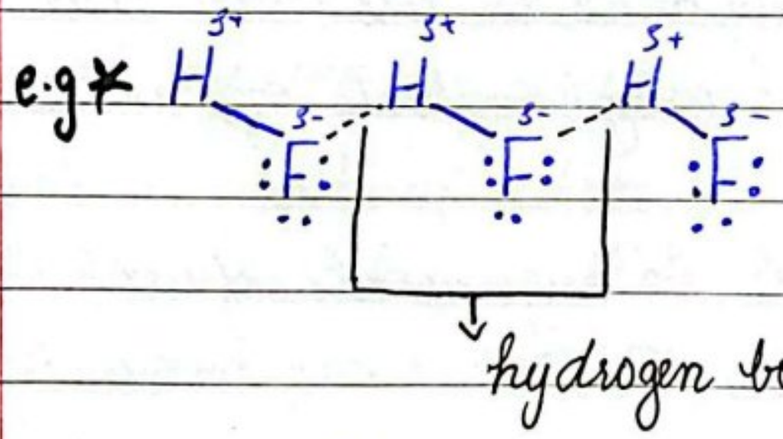
Date: Day:

3- Hydrogen Bonding:

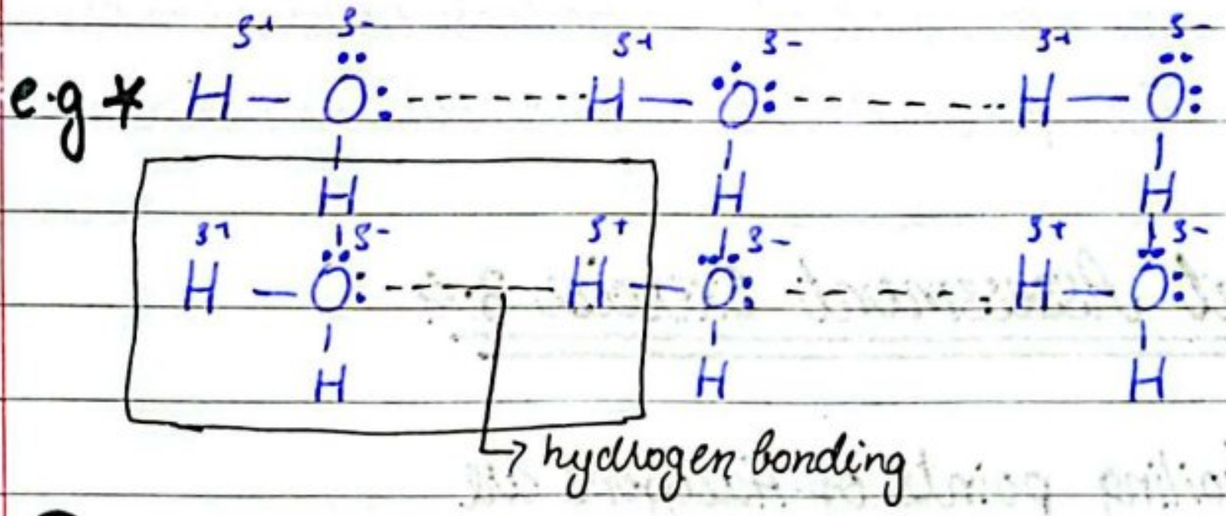
def: "Electrostatic force of attraction between 'H' atom of one molecule and lone pair of high E.N elements like O, N, F of other molecule is called hydrogen bonding."

- 20% effective
- It is a strong intermolecular force.
- Those covalent compounds which form hydrogen bonding with water molecules are water soluble.

e.g *



e.g *



Decreasing order of IMF:

Hydrogen bonding > dipole-dipole forces > L.D.F

PECULIAR BEHAVIOR OF HYDROGEN BONDING

Date: Day:

Ice is less dense than water:

Most of the solids are denser than liquid because in solid molecule are closely packed. This is not true for water. As in ice, there is a three-dimensional hydrogen bonded network of water molecule. This produces a rigid lattice in which oxygen atom is surrounded by a tetrahedron of hydrogen atom. In liquid the water molecules are further due to long hydrogen bond. Hence water in solid state occupy more space. So, the density of ice is less than that of liquid water. That is why ice floats on water.

Dependence: Melting and boiling points of covalent molecules depends upon IMF not on covalent bond in the molecules. Covalent bond is more strong. The forces between molecule is weaker.

Concept Assessment Exercise 3.4

1- a) The boiling points of Halogens are

Fluorine	-188°C	Chlorine	-35°C
Bromine	+59°C	Iodine	+184°C

Explain the trend in these boiling points

The boiling points of halogens increases down the group. Larger molecules have more electrons.

Fluorine = lowest

Chlorine = higher

Bromine = higher than chlorine

Iodine = highest than all.

Date: _____ Day: _____

b) The table lists the formulae and boiling points of some alkanes. Explain this trend.

Methane	CH_4	-164
Ethane	CH_3CH_3	-88
Propane	$CH_3CH_2CH_3$	-42
Butane	$CH_3CH_2CH_2CH_3$	0

As the Carbon atom increases the bigger the molecule is, more will be electrons and size will be larger. As we go down the B.P increases and the molecular mass increases.

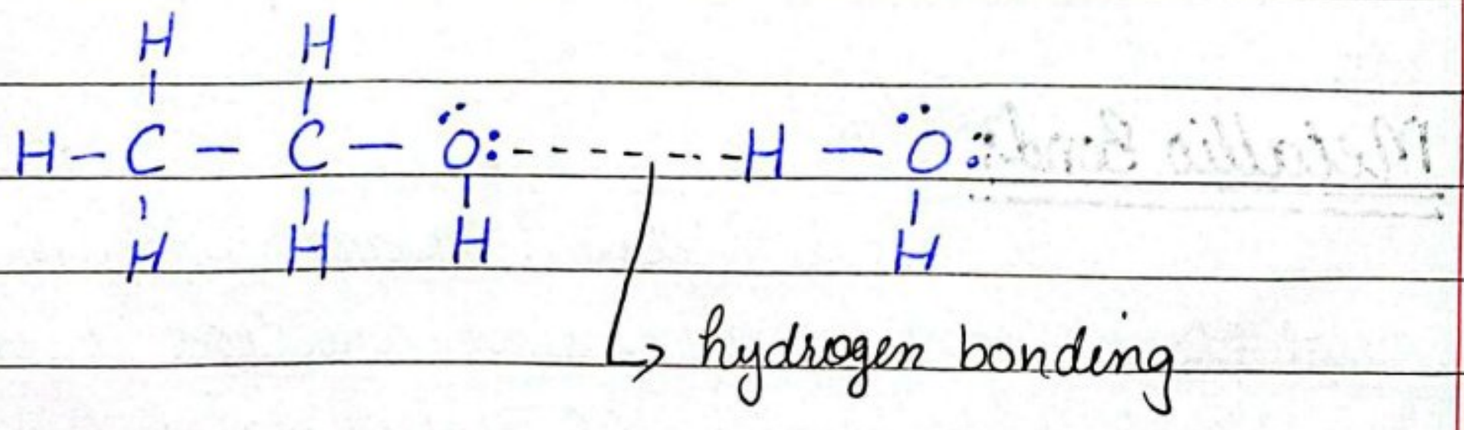
2- Br_2 and ICl have same number of electrons but boiling point of iodine monochloride is nearly 40°C higher than the boiling point of bromine. Explain breifly.

1) Bromine is nonpolar and has london dispersion forces that are weak forces and ICl has dipole dipole interaction have stronger interaction between their molecules.

2) ICl is polar because molecules are different and $\mu \neq 0$ whereas Br_2 has same molecules so $\mu = 0$.

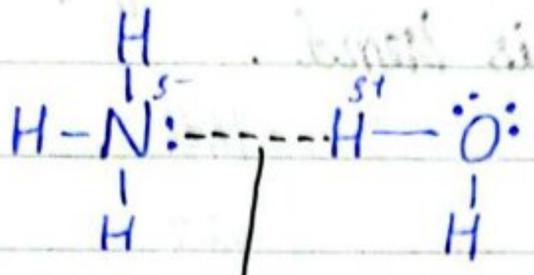
3- Draw diagram to show hydrogen bonding b/w the following molecules.

a) Ethanol and water

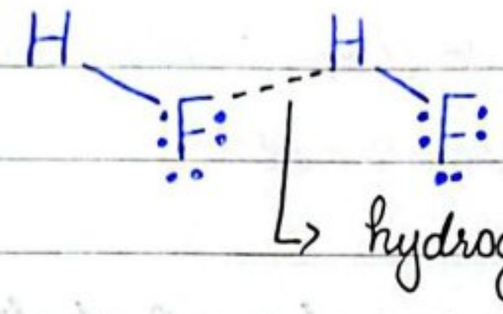


Date: Day:

b) Ammonia and water



c) two HF molecules



COMPARISON OF STRENGTH OF IONIC, COVALENT, METALLIC BONDING AND INTERMOLECUL FORCES

The strength of a bond depends on the specific elements involved.

Ionic Bond:

It contains strong electrostatic forces of attraction between positive and negative ion.

Metallic Bond:

It occurs between atoms within a metal, where electrons are shared and free to move through out the material. This creates sea of delocalized electrons that hold the metal atom together.

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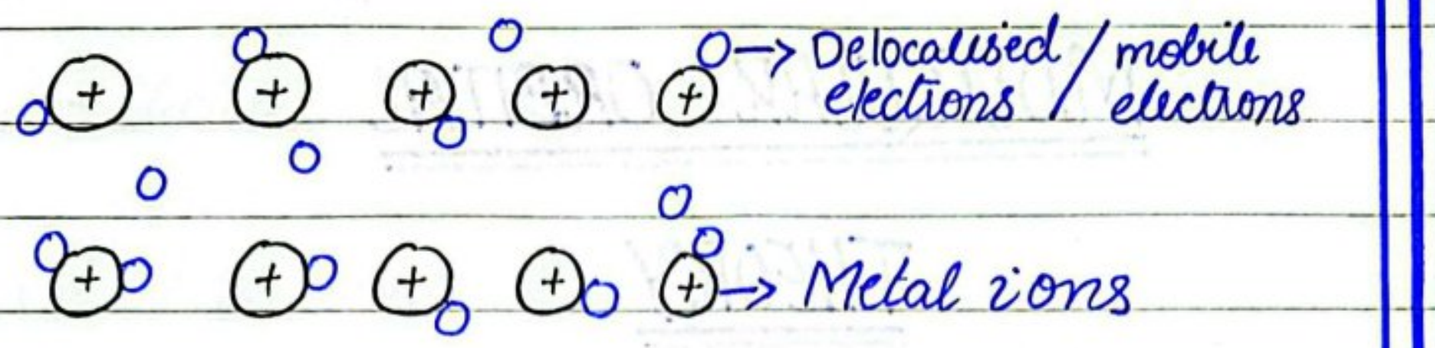


Electron Sea Theory

Covalent Bond:

It is formed by sharing of electron and results in formation of strong bond b/w atoms. The electrons are held in the region between the nuclei of the bonded atoms and the positive nuclei are attracted to these shared electrons. This leads to stable and strong bonding interaction.

Intermolecular forces: (IMF)

IMF are weaker forces than other and occurs between molecules, not within a single molecule. These forces are responsible for holding molecules together in a substance but are much weaker as compared to the forces within a molecule (covalent bond).

Substances	Bonding	M.P/°C	B.P/°C
Sodium Chloride	Ionic (E.F)	801	1465
Diamond	Giant Covalent (C.B)	4000	4830
Iron	Metallic (E.F)	1538	2862
Methane	Simple covalent (IMF)	-182	-161.6

Methane has lower M.P and B.P because IMF are weak. It requires less energy to break.

Date: Day:

MOLECULAR ORBITAL THEORY

Discovered By:

The M.O.T was proposed by Hund and Mullikan in 1932

Explanation:-

According to this theory, linear combination of atomic orbitals form new orbitals called molecular orbitals which are characteristics of the whole molecule. During this process identity of both atomic orbitals is lost. After the combenation of atomic orbitals different types of M.O are formed which differ in energy.

POSTULATES:

① Bonding and Antibonding Molecular Orbitals

- **Bonding Molecular Orbitals (B.M.O):**

  def:- Molecular orbital having energy less than atomic orbitals are called B.M.O

  - They are denoted by σ, π
  - In these orbitals electron density exist between two nuclei.
  - They are more stable.

- **Antibonding Molecular Orbitals (A.M.B.O):**

  def:- Molecular orbital having energy greater than atomic orbitals are called A.M.B.O.

  - They are denoted by $\overset{*}{\sigma}$, $\overset{*}{\pi}$

Date: Day:

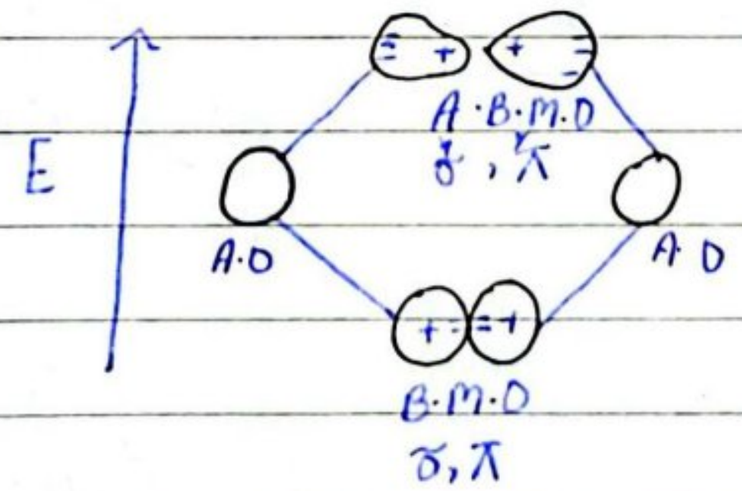
- Electron density exist away from the two nuclei.
- It is unstable.

*-> The filling of electrons occurs accordingly.
- Aufbau principle - Pauli's Exclusion Principle
- Hund's Rule

Types:

1- **Head on approach:** takes by the combination of s-s, s-p, p_x-p_x :orbitals - When "s" atomic orbitals overlap with each other two molecular orbitals σs of low energy and σ^*s of high energy are formed.

2- **Sideways approch:** takes place between p_y-p_y and p_z-p_z orbitals. When two p orbitals p_y or p_z atomic orbital overlap. $\pi p_y = \pi p_z$ of low energy and $\pi^* p_y = \pi^* p_y$ of high energy.



② Bond Order:

Total number of bonds formed when two atoms of atomic orbitals overlap with each other.

Date: Day:

$$\text{Bond order} = \frac{\text{No of e}^- \text{s in B.M.O} - \text{No's of e}^- \text{s in A.B.M.O}}{2}$$

* If bond order is zero, the molecule is unstable and it does not exsit.
* If bond order is not zero, the molecule is stable and it exsit.

③ Paramagnetic and Dimagnetic Substances :-

- **Paramagnetic Substance:**

  def:- "Substances which is attracted in magnetic field due to presence of unpaired electrons."

  e.g:- O_2.

- **Dimagnetic Substance:**

  def: "Substances which is weakely repelled in magnetic field due to presence of paired of electrons."

④ Relative energy of Molecular Orbitals:

- Energy level for F_2, O_2 and their +ve or -ve ions.

$$\sigma 1s < \overset{*}{\sigma} 1s < \sigma 2s < \overset{*}{\sigma} 2s < \sigma 2p_x < \pi 2p_y = \pi 2p_z < \overset{*}{\pi} 2p_y = \overset{*}{\pi} 2p_z < \overset{*}{\sigma} 2p_x$$

- Energy level for $N_2, C_2, Be_2, H_2, He_2$ and other.

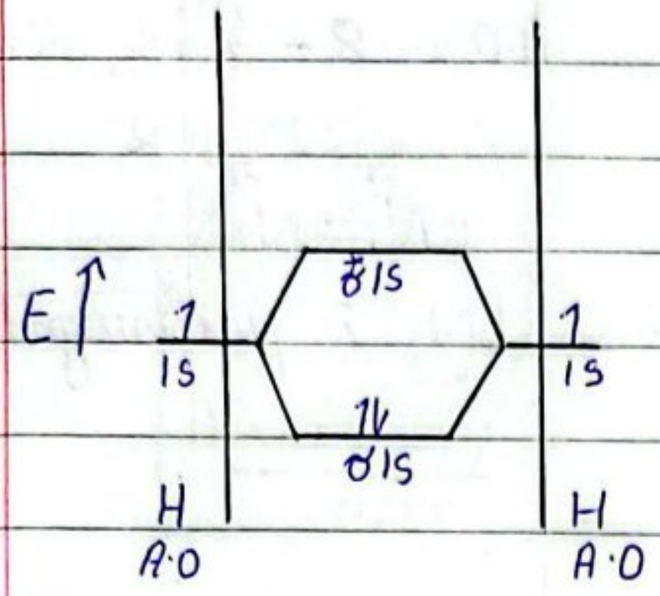
$$\sigma 1s < \overset{*}{\sigma} 1s < \sigma 2s < \overset{*}{\sigma} 2s < \pi 2p_y = \pi 2p_z < \sigma 2p_x < \overset{*}{\pi} 2p_y = \overset{*}{\pi} 2p_z < \overset{*}{\sigma} 2p_x.$$

Applications:

$H_2, He_2, N_2, O_2, F_2, Ne_2, N_2^{-2}, O_2^{+2}, F_2^{+2}$

Date: _____ Day: _____

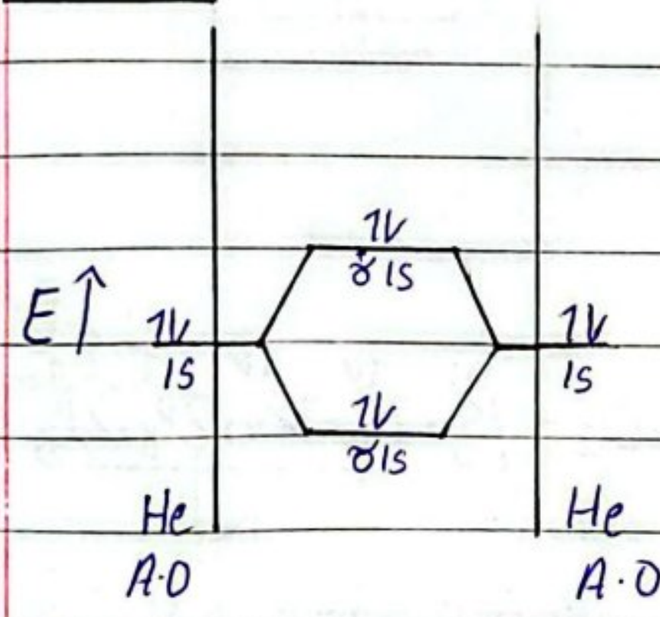
① H_2:-



Bond order = $\frac{\text{No of } e^- \text{ in B·MO} - \text{No's of } e^- \text{ in AB·MO}}{2}$

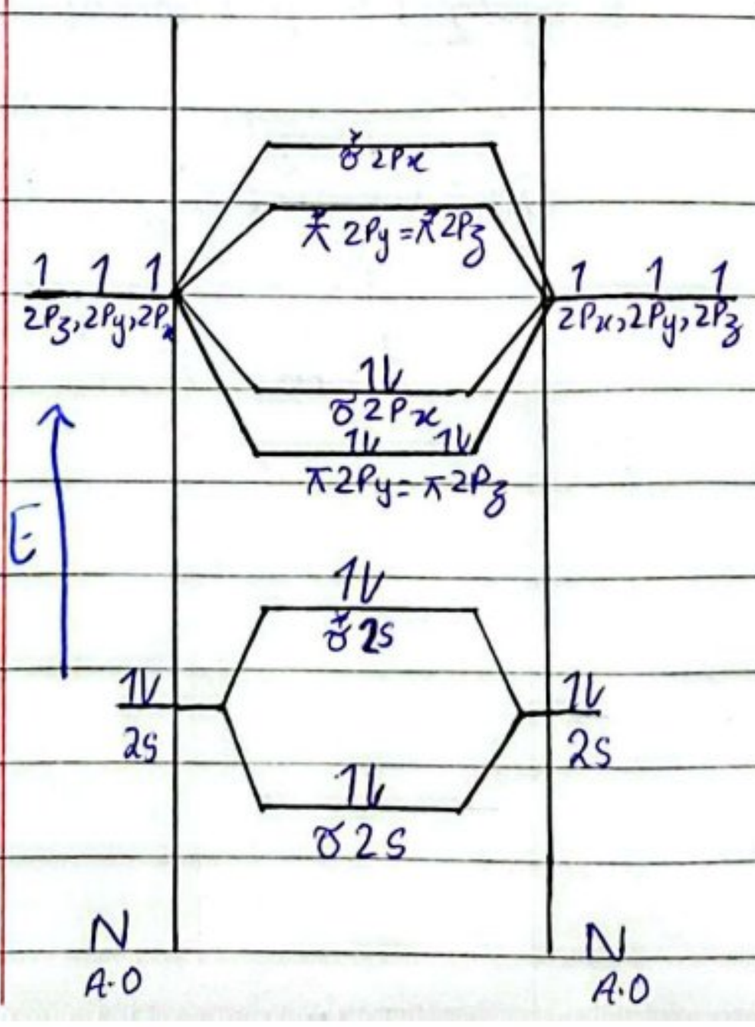
- $B.O = \frac{2-0}{2} = \frac{2}{2} = 1$
- H—H
- H_2 is diamagnetic
- $H_2 = \sigma 1S^{1v}$

② He_2:-



- $B.O = \frac{2-2}{2} = \frac{0}{2} = 0$
- He_2 molecule doesnot exist

③ N_2:-



- $N_2 = N_7 = \underbrace{1s^{1v}}_{K}, \underbrace{2s^{1v}, 2P_x^{1}, 2P_y^{1}, 2P_z^{1}}_{L}$
- $B.O = \frac{8-2}{2} = \frac{6}{2} = 3$
- $N \equiv N$
- N_2 is diamagnetic
- $N_2 = \sigma 1s^{1v}, \sigma^* 1s^{1v}, \sigma 2s^{1v}, \sigma^* 2s^{1v}, \pi 2P_y^{1v} = \pi 2P_z^{1v}, \sigma 2P_x^{1v}$

Date: Day:

④ N_2^{-2} :-

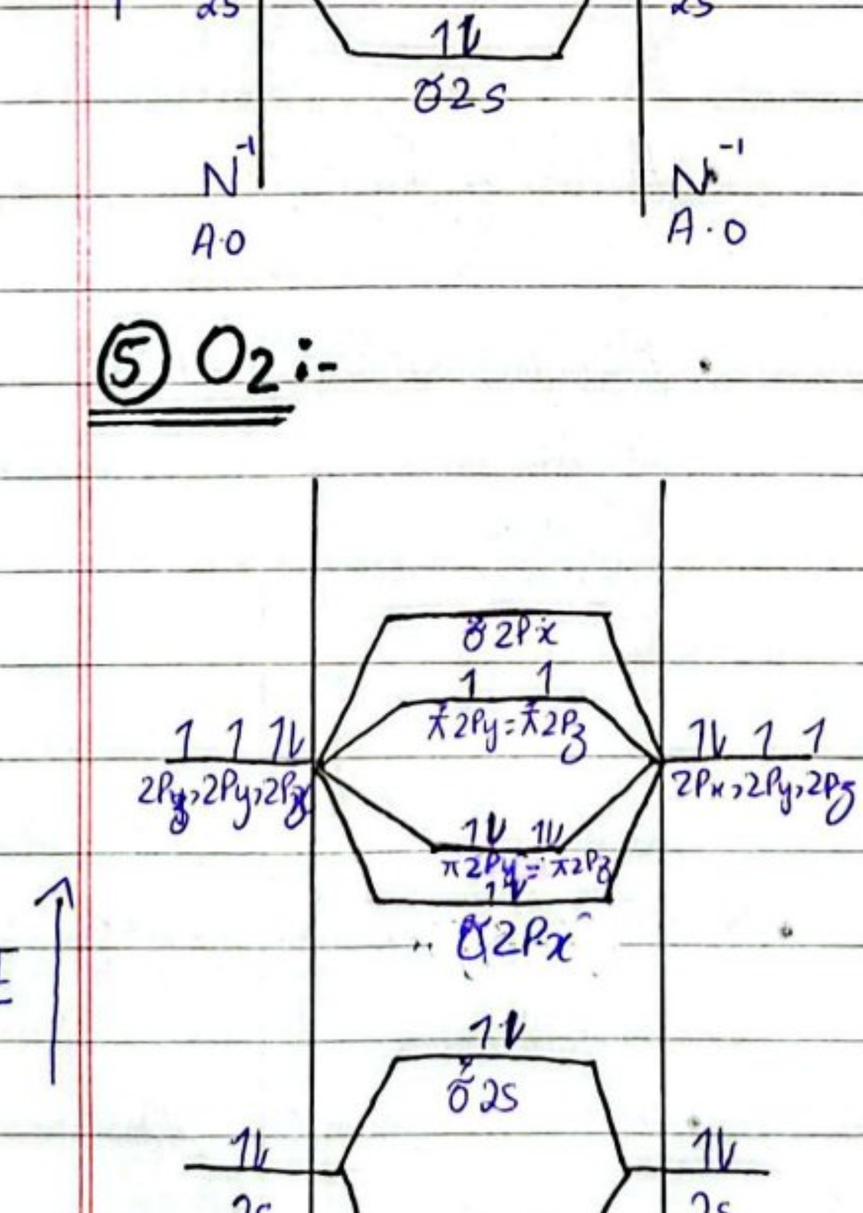
$N_2^{-2} = \underline{1S^{2}}, \underline{2S^{2}, 2P_x^{2}, 2P_y^{1}, 2P_z^{1}}$

K L

$B.O = \frac{8-4}{2} = \frac{4}{2} = 2$

$N^{+} = N^{-1}$

N_2 is paramagnatic



⑤ O_2 :-

$O_2 = \underline{1S^{2}}, \underline{2S^{2}, 2P_x^{2}, 2P_y^{2}, 2P_z^{2}}$

K L

$B.O = \frac{8-4}{2} = \frac{4}{2} = 2$

$O = O$

O_2 is paramagnetic

$O_2 = \sigma 1s^{2}, \sigma^{*} 1s^{2}, \sigma 2s^{2}, \sigma^{*} 2s^{2}$

$\sigma 2P_x^{2}, \pi 2P_y = \pi 2P_z$

$\pi^{*} 2P_y^{1} = \pi^{*} 2P_z^{1}$

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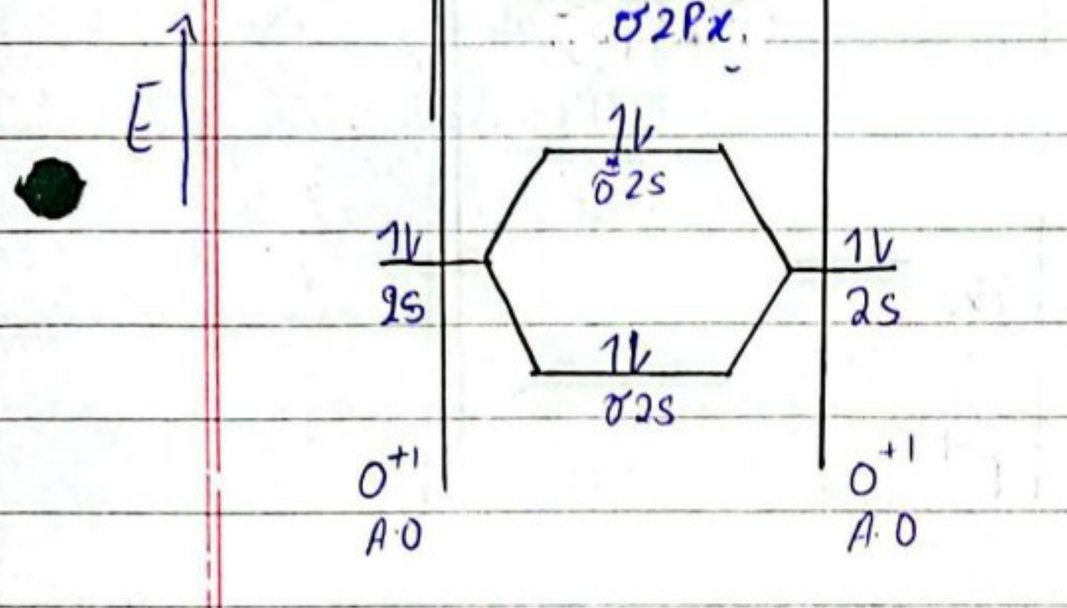
⑥ O_2^{+2}

$O_2^{+2} = 1s^{2}, 2s^{2}, 2P_x^{2}, 2P_y^{1}, 2P_z^{1}$

$B.O = \frac{8-2}{2} = \frac{6}{2} = 3$

$O^{+1} \equiv O^{+1}$

O_2^{+2} is diamagnetic.



$O_2^{+2} = \tilde{\sigma}1s^{2}, \tilde{\sigma}1s^{2}, \tilde{\sigma}2s^{2}, \tilde{\sigma}2s^{2}$
$\tilde{\sigma}2P_x^{2}, \pi 2P_y^{2} = \pi 2P_z^{2},$

⑦ F_2 :-

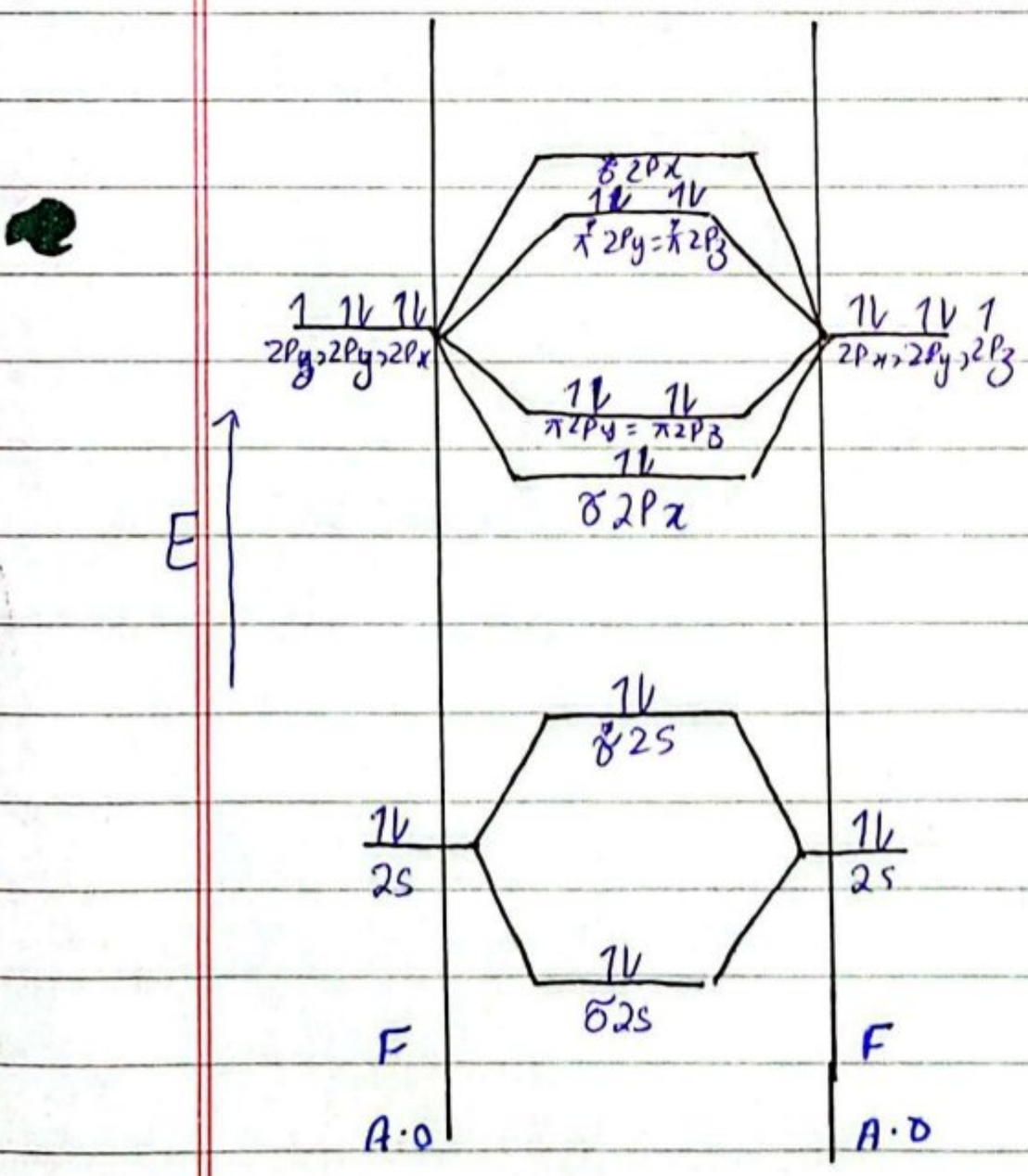
$F_2 = \underbrace{1s^{2}}_{K}, \underbrace{2s^{2}, 2P_x^{2}, 2P_y^{2}, 2P_z^{1}}_{L}$

$B.O = \frac{8-6}{2} = \frac{2}{2} = 1$

F–F

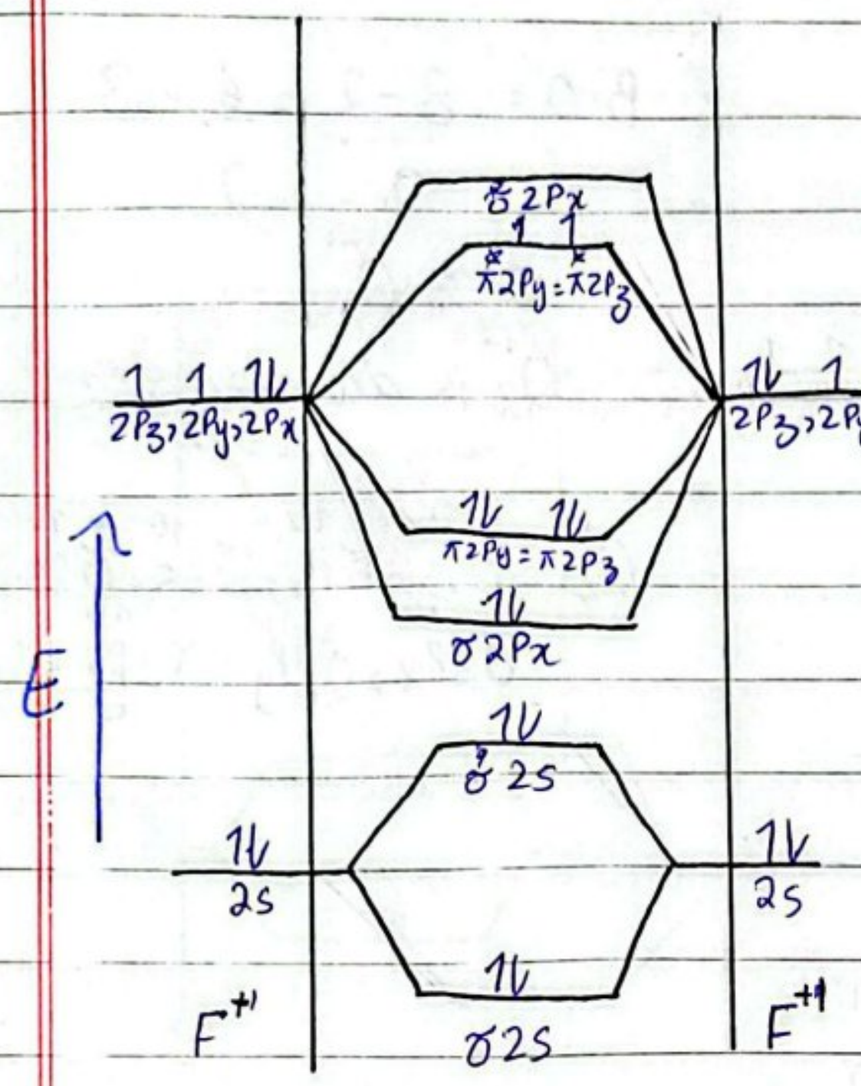
F_2 is a diamagnetic

$F_2 = \sigma 1s^{2}, \tilde{\sigma}1s^{2}, \tilde{\sigma}2s^{2}, \tilde{\sigma}2s^{2}$
$\tilde{\sigma}2P_x^{2}, \pi 2P_y^{2} = \pi 2P_z^{2},$
$\pi 2P_y^{2} = \pi 2P_z^{2}, \tilde{\sigma}2P_x^{2}$



Date: Day:

⑧ F_2^{+2} :-



$F_2^{-1} = \overset{1v}{1s}, \overset{1v}{2s}, \overset{1v}{2P_x}, \overset{1}{2P_y}, \overset{1}{2P_z}$

K L

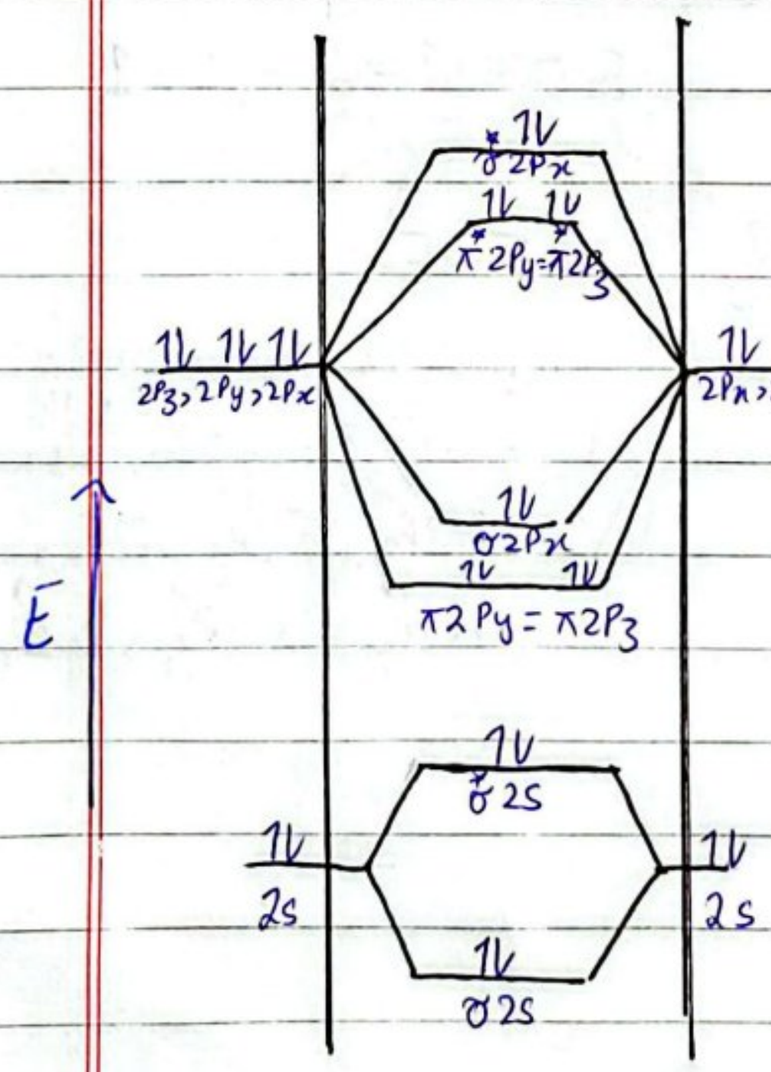
$B.O = \frac{8-4}{2} = \frac{4}{2} = 2$

$F^{-1} = F^{-1}$

F_2^{+2} is paramagnetic

$F_2^{-2} = \overset{1v}{\sigma 1s}, \overset{1v}{\sigma^* 1s}, \overset{1v}{\sigma 2s}, \overset{1v}{\sigma^* 2s}, \overset{1v}{\sigma 2P_x}, \overset{1v}{\pi 2P_y = \pi 2P_z}, \overset{1v}{\pi^* 2P_y = \pi^* 2P_z}, \overset{1v}{\sigma^* 2P_x}$

⑨ Ne_2



$Ne_2 = \overset{1v}{1s}, \overset{1v}{2s}, \overset{1v}{2P_x}, \overset{1v}{2P_y}, \overset{1v}{2P_z}$

k L

$B.O = \frac{8-8}{2} = \frac{0}{2} = 0$

Ne_2 does not exist.

Date: _____ Day: _____

EXERCISE

1. Choose the correct Answer.

(i) c (ii) d (iii) d
(iv) b (v) d (vi) d
(vii) c (viii) d (ix) d
(x) b (xi) c (xii) b
(xiii) d

2. The table shows the atomic number and boiling points of some noble gases.

Gas	helium	neon	argon	krypton	Xenon
Atomic no's	-12	-110	-118	-136	-154
Boiling Point	-253	-246	-186	-152	-107

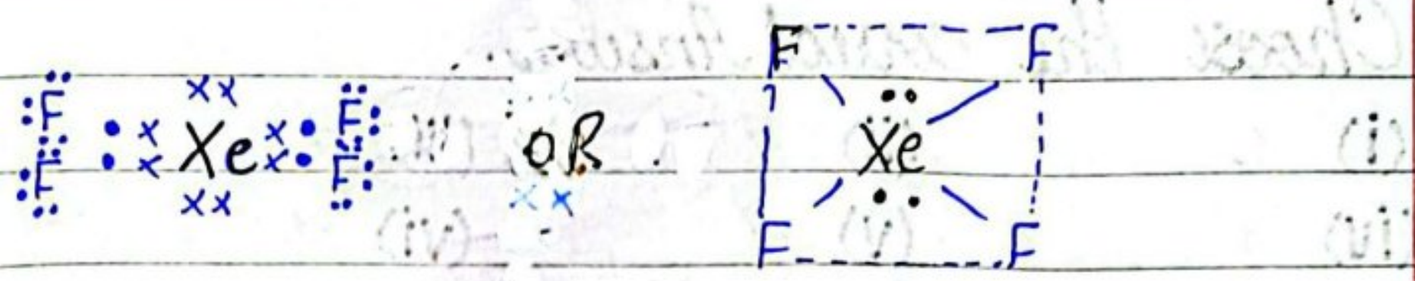
a) Explain this trend in boiling points.
down the group the B.P increases with atomic number because size is larger, greater will be intermolecular forces and number of electron increases.

b) Xenon forms a number of covalently bonded compounds with fluorine.
Valence electron of xenon = 8
Valence electron of Fluorine = 7

→ → This bond will form 4 single bonds between xenon and fluorine atom.

Date: _____ Day: _____

c) draw a dot and cross diagram for XeF_4.



d) Suggest the shape for XeF_4. Explain why you choose this shape.

Shape: Square Planar

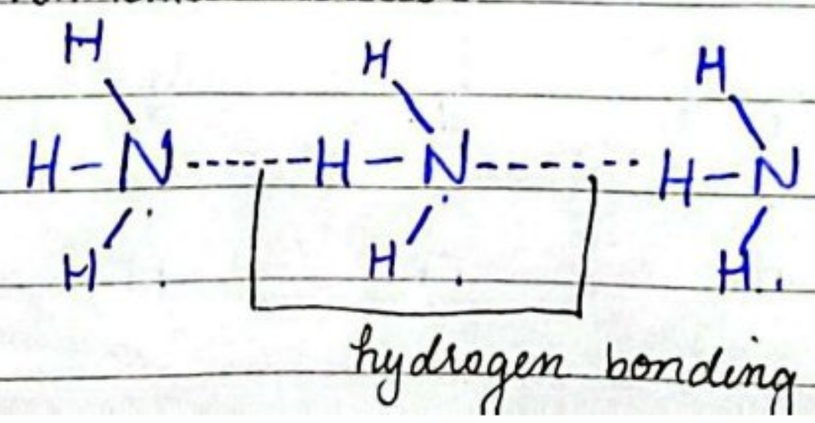
Angle: 90°

3- Aluminium Chloride, $AlCl_3$ and ammonia NH_3, are both covalent molecules.

a) Draw a diagram of an ammonia molecule, showing its shape. Show any lone pairs of electrons. Also state the bond angle HNH in in the ammia molecule.

- Shape: Trigonal pyramid
- lone pairs: 1
- Bond angle: <109.5, 107.5

b) What type of forces are present in ammonia molecule - Draw diagram to show forces blw Ammonia molecule.



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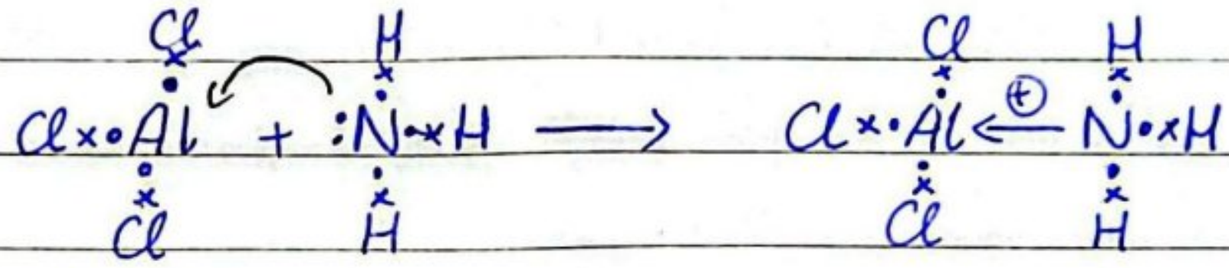
c) an ammonia molecule and an aluminium chloride molecule can join together by forming a co-ordinate bond.

i) Explain how co-ordinate bond is formed.

def: It is formed by sharing of electrons or shared pair of electron and denoted by one of bond atom.

- In NH_3 and $AlCl_3$ both forme co-ordinate bond when NH_3 share its pair of election with $AlCl_3$.

(ii) Dot and cross diagram to show bonding in NH_3 and $AlCl_3$, H_3NAlCl_3



Alleminium chloride — Ammonium Molecule — H_3NAlCl_3

4. Electronegativity values can be used to predict the polarity of bonds.

a) Explain the term Electronegativity.

(Answers on Page 1)

b) The electronegativity values for 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 bonds by coping the bonded atom below and adding $\delta+$ or $\delta-$ above each atom.

1. $\overset{\delta+}{H}-\overset{\delta-}{I}$ => 2.5 - 2.1 = 0.4 (polar)
2. $\overset{\delta-}{F}-\overset{\delta+}{I}$ => 4.0 = 2.5 = 1.5 (polar)
3. $C-Cl$ => 3.0 - 2.5 = 0.5 (polar)

Higher E.N ; ① Iodine ② Fluorine ③ Chlorine

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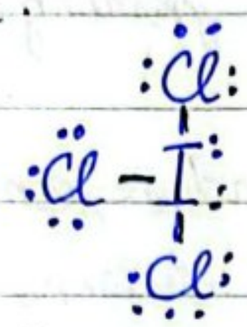
c) Describe the shape of this ICl_3 molecule. Also mention bond angle.



Bond angle = 90°

d) The boiling points of hydrogen halides are

Hydrogen Halide	HF	HCl	HBr	HI
Boiling point/°C	+20	-85	-67	-35

i) Explain the trend in B.P from HCl to HI

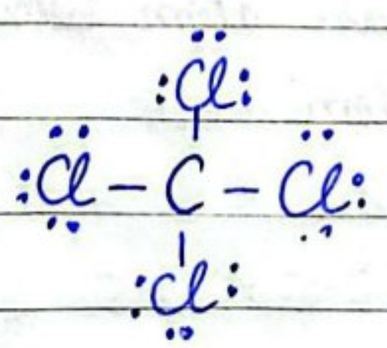
From HCl to HI the B.P increases as number of electrons increase, size will be greater and IMF will also increases.

ii) Why B.P of HF is so much high than the HCl.

Because HF has strong hydrogen bonding due to which more energy is required to break bond whereas in HCl there are weaker dipole dipole interaction.

e) Tetrachloromethane, CCl_4 is a non-polar molecule. Draw a diagram to show the shape of this molecule. Explain why this molecule is non-polar.

C = 2.5 Cl = 3.0 $\Rightarrow$ 3.0 - 2.5 = 0.5



shape: Tetrahydral.

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This molecule is non-polar because is tetrahedral structure. It $\mu = 0$ as atoms are same. dipole of bond exert equal and opposite effect so it cancel the charge.

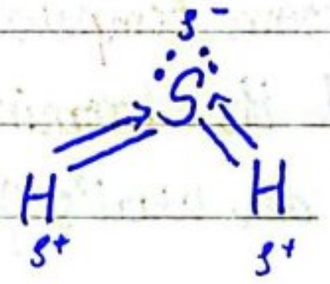
5. a. H_2S is a covalent compound. Explain the type of hybridization. also write bond angle in HSH. Also show on your diagram the partial charges on each atom as $\delta+$ and $\delta-$ and an arrow showing the exact direction of the dipole in the molecule as a whole.

Sp^3-Hybridization:-

"Mixing of one s and 3p that is P_x, P_y, P_z. forming 4 new orbitals of equal energy and same shape is called sp^3 hybridization."

. Bondangle: <109.5 or 104.5

Diagram:



b- Oxygen, Sulphur, Selenium are in the same group in the periodic Table.

i Explain why hydrogen selenide, H_2Se has a higher boiling point than hydrogen sulphide, H_2S.

H_2Se has higher boiling point as down the group the number of electron increases, their size increases, greater will be the London dispersion forces.

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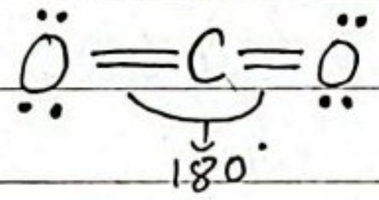
ii) Explain why the Boiling point of water is so much higher than the boiling point of Hydrogen Sulphide.

Boiling Point of H_2O is higher because of presence of strong hydrogen bonds between their molecule. This force requires more energy to break. Whereas H_2S have dipole-dipole interaction. As compare to hydrogen bond, dipole-dipole interaction are weaker then it so water has high B.P.

6. a. Describe the shape of the CO_2 molecule.

In CO_2 the shape is linear. because of two double bonds present in it.



Bond angle = 180°

b. Bromine is a liquid at room temperature. Weak van der Waal's forces hold the bromine molecules together. Describe how van der Waal's forces arises.

Van der waal's forces arises from temporary dipoles because of uneven electron distribution due to the constant movement of electrons. This creates weak attractions between them.

7. Water is extensively hydrogen bonded. - This gives it anomalous properties.

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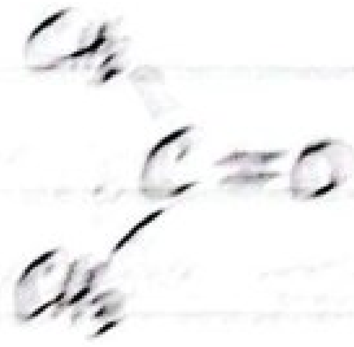
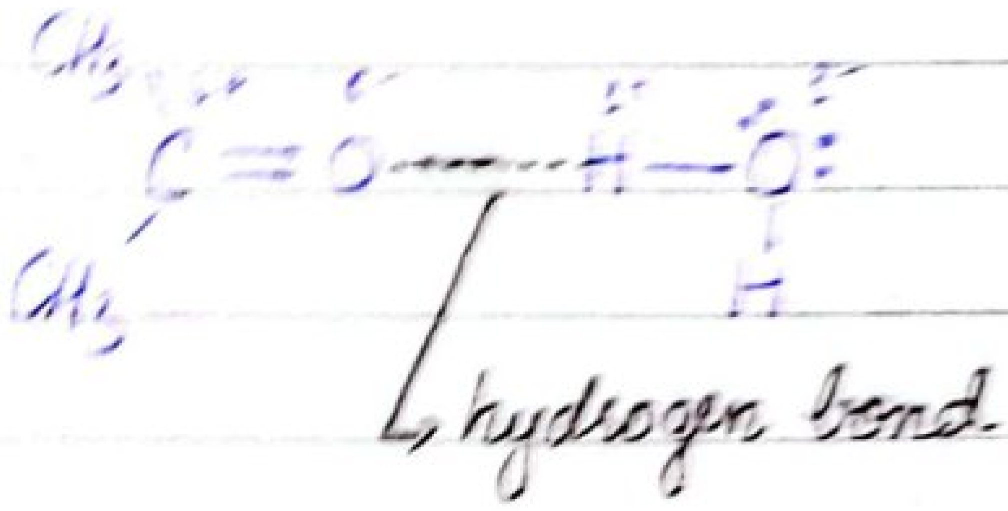
1. Explain why ice is less dense than liquid water. Also state two other anomalous properties of water.

(Answer on Page 32)

2. Propanone has structure below.



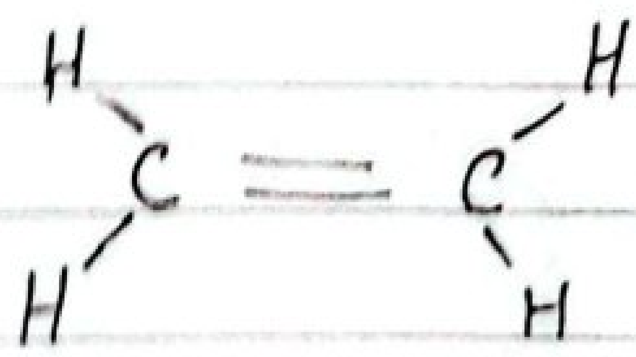
When propanone dissolves in water, it forms a hydrogen bond with water. Draw a diagram to show a propanone molecule and a water molecule forming a hydrogen bond.



C. (i) Propanone has a double bond. One of the bonds is a σ bond, other is pi bond. Explain difference b/w π and σ bond in terms of how they are formed.

(Answer on Page 20 + 21)

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



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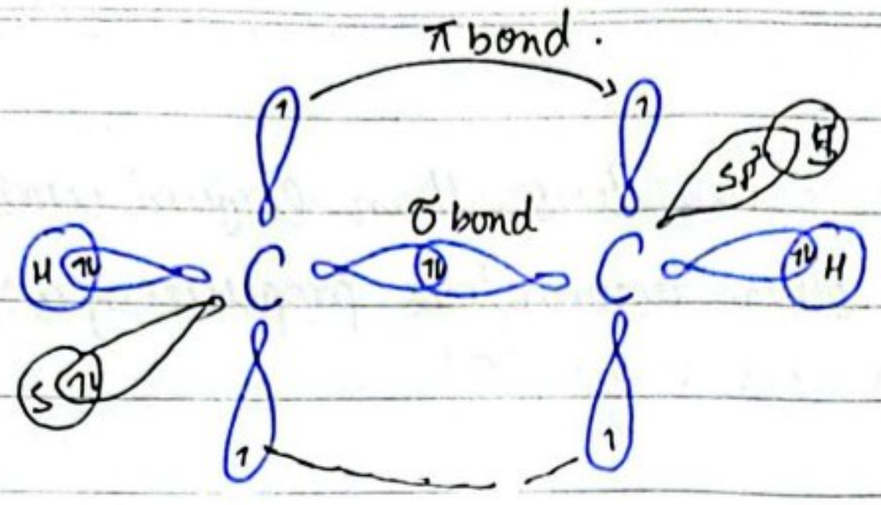

8- Energies of orbitals can be explained by molecular orbital Theory. It has been observed that in case of Nitrogen molecule $\sigma 2P_x$ is higher in energy than $\pi 2P_y$ and $\pi 2P_z$.

a. Draw molecular orbitals energy diagram for nitrogen molecule.

(Answer On Page 45)

b. Give reason why the $\sigma 2P_x$ energy is greater than $\pi 2P_y$ and $\pi 2P_z$.

As $\sigma 2P_x$ orbital forms by end to end overlap and $\pi 2P_y$ and $\pi 2P_z$ forms by side by side overlap due to which $\sigma 2P_x$ is stronger and $\pi 2P_y$ and $\pi 2P_z$ is weak.

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 molecule on basis of hybridization with the help of diogram.

(Answer on Page 25)

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10. Explain the magnetic properties of O_2, O_2^+ and N_2 by applying M.O.T.

(Answer on Page 45, 46, 47)

11. Differenciate between a sigma bond and a pi-bond.

(Answer on Page 20+21