

CHEMICAL BONDING & MOLECULAR STRUCTURE



Solubility of ionic bond:

- > Solvation of Hydration:
 - Whenever any compound generally ionic or polar covalent is dissolved in a polar solvent or in water, then different ions of the compound will get separated and will get surrounded by polar solvent molecules. This process is known as solvation or hydration. Energy released in this process is known as solvation energy or hydration energy.

This ionic compound will be soluble only if solvation energy (S.E.) is more than the lattice energy.





> Factors affecting solvation energy of hydration energy

S.E $\propto Z_+ Z_ \propto \left(\frac{1}{r_+} + \frac{1}{r_-}\right)$



$$\propto \left(1 - \frac{1}{\epsilon_r}\right)$$
 (nature of solvent) where ϵ_r is dielectric constant

Greater the polarity. Greater will be ϵ_r

$$\epsilon_r \uparrow \Rightarrow \frac{1}{\epsilon_r} \downarrow, \quad 1 - \frac{1}{\epsilon_r} \uparrow \Rightarrow S.E.\uparrow$$

 $H_2 0 \quad CH_3 OH \quad C_2 H_5 OH \quad C_6 H_6$
 $\epsilon_r \quad 81 \quad 34 \quad 27 \quad 34$

Applications of hydration energy:

a) Size of the hydrated ions: Greater the hydration of the ion, grater will be its hydrated radii $Li^+(aq) > Na^+(aq)$



b) Mobility of the ion: More is the hydration, smaller will be the mobility of the ions. $Li^+(aq) < Na^+(aq) < K^+(aq) < Rb^+(aq) < Cs^+(aq)$

c) Electrical conductance: It is related to mobility so it follows the same order.

How to decide solubility?

Now we have
$$\rightarrow$$
 L.E $\propto \frac{1}{r_+ + r_-}$ and S.E. $\propto \left(\frac{1}{r_+} + \frac{1}{r_-}\right)$

From these 2 equations, it can be mathematically proved that if the difference between the radii of cation and anion is large, then solvation energy will dominate and if radii are comparable or difference is small, then lattice energy will dominate.

Greater the difference between radii, greater will be solubility.





Decrement in solvation energy > decrement in lattice energy.

Solubility order:

Increase difference in radii ≠ solubility ≠ and vice versa a) LiF → LiI increase

- b) NaF \rightarrow NaI increase
- c) $LiF \rightarrow CsF$ increase
- d) $CsF \rightarrow CsI$ decrease



- e) LiI \rightarrow CsI decrease
- f) $Be(OH)_2 \rightarrow Be(OH)_2$ increase
- g) LiOH \rightarrow CsOH increase
- h) $BeSO_4 \rightarrow BeSO_4$ decrease white ppt
- i) $LiClO_4 \rightarrow CsClO_4$ decrease
- j) $BeF_2 \rightarrow BeI_2$ decrease
- k) $AgF \rightarrow AgI$ decrease



Facts about solubility in H_2O :

- a) All the salts of NH_4^+ . Alkali metals are soluble except LiF, Li_3PO_4 , $Li_2C_2O_4$ and Li_2CO_3 .
- b) All the NO_3^- , ClO_4^- , and CH_3COO^- are soluble.
- c) S^{2-} (Sulphides) of NH_4^+ alkali metals or alkaline earth metal are soluble, whereas
- d) Sulphates of all elements are soluble except



- e) Hydroxides of NH_4^+ , alkali metals and $Ba(OH)_2$ are soluble, whereas rest are insoluble.
- f) Halides are generally soluble except AgCl (white), AgBr (pale yellow), AgI (yellow), and *PbI*₂ (yellow).



Prediction of hybridized state of central atom in a molecule:

| H = $\frac{1}{V}$ = No. of +No. of electron in monovale valence shell of surroundin central atom atoms (X) | - Charge nt on cation ng (C) | + Charge on Anion (A) |
|---|------------------------------------|--------------------------|
|---|------------------------------------|--------------------------|

| Valu e of H | 2 | 3 | 4 | 5 | 6 | 7 |
|-----------------------|----|--------|------------------------|-------------------|--------------------------------|--------------------------------|
| Hybr idiza tion | sp | sp^2 | <i>sp</i> ³ | sp ³ d | sp ³ d ² | sp ³ d ³ |



Examples:

a) Type I

Central atom is surrounded by monovalent atoms only BeF₂, BCl₄, NCl₃, PCl₃, PCl₅, NH₃, OF₂, H₂O, IF₇, CIF₃, SF₆, XeF₄ For examples

CCl₄ H =
$$\frac{1}{2}$$
 (4 + 4 − 0 + 0)
Surrounding monovalent
No. of electron in valency shell of 'C' H=4 ⇒
C is sp³ hybridized

 \Rightarrow l.p=H –b.p.= 4-4 =0



b) Type II

Central atom is surrounded by divalent atoms only

 $SO_2, SO_3, CO_2, CS_2, XeO_3$

For examples

$$XeO_3 \quad H = \frac{1}{2} \quad (8 + 0 - 0 + 0) = 4$$

$$As \text{ oxygen is divalent}$$

$$As oxygen is divalent$$

$$No. of electron in valency shell of 'Xe' H=4 \Rightarrow$$

$$C \text{ is } sp^3 \text{ hyb.}$$

 \Rightarrow l.p=H –b.p.= 4-3 =1



c) Type III

Central atom is surrounded by monovalent as well as divalent atoms only

 $COCl_2, XeO_2F_2. POCl_3$

For examples

POCl₃ H =
$$\frac{1}{2}$$
 (5 + 3 - 0 + 0)

⇒b.p=4

H= 4 \Rightarrow 'Xe' is Sp^3 hybridization

'O' in not inclined as it is divalent \Rightarrow l.p=H –b.p.= 4-4 =0



d) Type IV

Hybridization in anions

 SO_4^{2-} , , CO_3^{2-} , PO_4^{3-} , NO_2^{-} , NO_3^{-}

For example, NO_2^-

$$\Rightarrow$$
 bp = 2

 $H = \frac{1}{2}(5 + 0 - 0 + 1)$ from -ve charge $\Rightarrow Ip = 3 - 2 = 1$

 $H=3 \Rightarrow N \text{ is sp}^2 \text{ hybridized}$



e) Type V

Hybridization in cations H_3O^+ and NH_4^+

For example, H_3O^+ \Rightarrow bp = 2

 $H = \frac{1}{2}(6 + 3 - 1 + 0) = 4$ charge of cation

 \Rightarrow 0 is sp³ hybridized

 \Rightarrow bp = 3 Ip = 4-3 = 1



f) Type VI

Hybridization in complex ions ICI, I_3^-

For example, $I_3^- \Rightarrow I$ $I_2^- \Rightarrow$ one I is central

 $H = \frac{1}{2}(7 + 2 - 0 + 1) = 5I$ is sp³d hybridized.

bp = 2

Ip = 5-2 =3.



Dipole moment:

| ioment: | Formula | Molecular Geometry | Dipole Moment |
|---------|-----------------|---|--|
| | AX | Linear | can be non – zero |
| | AX ² | Linear Bent | can be non – zero |
| | AX ₃ | Trigonal planar Trigonal pyramidal T - shaped | zero can be non – zero can be non – zero |
| | AX ₄ | Tetrahedral Square planar | zero zero can be non – |



Applications of Dipole Moment

Applications of dipole moment are as follows:

To Predict Nature of Molecule

Molecules with non-zero dipole moments $(\mu \neq 0)$ are called polar molecules, where as those with zero dipole moments $(\mu \neq 0)$ are called non – polar molecules.

In preceding the polarity of bonds: higher the dipole moment, greater is the polar character.

For example,



| H - F | 1.91D | 45% ionic character (highly polar bond) |
|--------|-------------|---|
| H – CI | 1.03D | 17% ionic character |
| H – Br | 0.8D | 11% ionic character |
| H - I | 0.4D | 4% ionic character (least polar bond) |

To Determine Shape of Molecules

If $\mu_{resultant} = 0 \Rightarrow$ regular geometry

If $\mu_{resultant} \neq 0 \Rightarrow$ irregular geometry



To Determine % Ionic Character.

The percentage ionic character of a polar covalent bond can be calculated as follows:

Percentage ionic character = <u>Observed dipole moment</u> × 100 <u>dipole moment for 100% ionic bond</u> × 100

$$=\frac{\mu_{obs}}{\mu_{ionic}}\times 100$$



MOLECULAR ORBITAL THEORY



MOLECULAR ORBITAL THEORY:

➤ This theory was developed by F. Hund and R.S. Mullikan.



Salient Features of MO Theory.

The atomic orbitals of comparable energies and proper symmetry combine to form molecular orbitals.



Sailent Features of MO Theory.

- An electron in an atomic orbital is influenced by one nucleus while in a molecular orbital, electron is influenced by two (or) more nuclei depending upon the no.of atoms invloved in the formation of atomic orbital.
- ➤ The atomic orbital is monocentric, while a molecular orbital is polycentric.
- When two atomic orbitals combine, two molecular orbitals are formed.
- ➤ The no.of molecular orbitals formed is equal to the no.of combining atomic orbitals.



Salient Features of MO Theory.

One is bonding molecular orbital, While the other is called anti bonding molecular orbital.





Salient Features of MO Theory.

- ➤ The bonding molecular orbital has lower energy and hence greater stability.
- ➤ The anti bonding molecular orbital has higher energy and hence lesser stability.
- The electron probability distribution around a nucleus in an atom is given by an atomic orbital.
- Like atomic orbitals, the molecular orbitals are filled in accordance with the 1) Aufbau 2) Pauli's & 3) Hund's rules.



Salient Features of MO Theory.

- Linear combination of atomic orbitals [LCAO].
- **It is applicable to homonuclear diatomic molecule like H**_{2.}
- ➢ Hydrogen which consists of 2 H atoms A&B.
- **Each hydrogen atom in the ground state has 1e⁻ in '1s' orbital.**
- > The atomic orbital's of these atoms may be represented by the wave functions $\psi_A \& \psi_B$.

 $\psi_{MO} = \psi_A + \psi_B$.

 $\psi_{MO} = \psi_A - \psi_B$.



> The molecular orbitals are represented by σ , σ^* and π, π^*

 $\sigma = \psi_A + \psi_B$ \longrightarrow Bonding molecular orbital.

 $\sigma^* = \psi_A - \psi_B \longrightarrow$ Antibonding molecular orbital.

> Constructive interference



- > They have less repulsion in between the nuclei.
- > Destructive interference.











1. Molecular orbitals filled according to the _____ principles.

a) Aufbau

b) Pauli's

c) Hund's





2. What are molecular orbitals?

a) σ

b) σ*

c) π, π*





CONDITIONS FOR THE COMBINATION OF ATOMIC ORBITALS



CONDITIONS FOR THE COMBINATION OF ATOMIC ORBITAL'S:

The combining atomic orbitals must have the same energy (or) nearly the same energy.

Ex:

- '1s' orbital can combine with another '1s' orbital but not '2s' orbital, because energy of '2s' orbital is appreciably high.
- The combining atomic orbitals must have the same symmetry about the molecular axis.



Ex:

- > '2P_z' orbital of one atom can combine with '2P_z' orbital of the other atom but not with $2P_x(or) 2P_y$ orbitals because of their different symmetries.
- > The combining atomic orbitals must overlap to the maximum extent.
- ➢ Greater the extent of overlap, the greater will be the electron-density between the nuclei of a molecular orbital.



ENERGY LEVEL DIAGRAM FOR MOLECULAR ORBITAL:

Upto 14e⁻

$$\sigma_{1s} < \sigma *_{1s} < \sigma_{2s} < \sigma *_{2s} < (\pi_{2Px} = \pi_{2Py}) < \sigma_{2Pz} < (\pi *_{2Px} = \pi *_{2Py}) < \sigma *_{2Pz}$$

Above 14e⁻

 $\sigma_{1s} < \sigma^{*}_{1s} < \sigma_{2s} < \sigma^{*}_{2s} < \sigma_{2Pz} < (\pi_{2Px} = \pi_{2Py}) < (\pi^{*}_{2Px} = \pi^{*}_{2Py}) < \sigma^{*}_{2Pz}$



ELECTRONIC CONFIGURATION AND MOLECULAR BEHAVIOUR

➤ The distribution of electrons among various molecular orbitals is called electronic configuration.

STABILITY OF THE MOLECULE

- > N_b = Number of electrons in bonding orbitals.
- \triangleright N_a = Number of electrons in anti bonding orbitals.

CASE-1: If $N_b > N_a$, The molecule is stable.

CASE-2: If $N_b < N_a$, The molecule is unstable.



- **1.** Which of the following is incorrect statement about the combination of atomic orbitals...
 - a) Combining atomic orbitals must have the same symmetry
 - $\frac{1}{2}$ 2p-orbital of Ne atom can combine with 2p_z of other
 - c) 1s orbital can't combine with 2s-orbital
 - d) all the above


MOED Of N₂, O₂ AND BOND ORDER



Examples of molecular orbital

Explain the formation of N₂ on the basis of MOT

Ans : N₂ molecule is formed by the combination of two N atoms.

N = 7 electrons

 \therefore N₂ = 14 electrons





MOLECULAR ORBITAL DIAGRAM

Bond order
$$=\frac{Nb - Na}{2} = \frac{10 - 4}{2} = 3$$

Magnetic property = Diamagnetic

E.C. =KK
$$(\sigma 2s)^2(\sigma^* 2s)^2(\pi 2p_x^2)$$

 $(\pi 2p_y^2)$
 $(\sigma 2p_z)^2$



- > There is a double bond between two oxygen atoms.
- > Due to the presence of unpaired electrons it is paramagnetic nature.
- > In O₂ molecule $1\sigma \& 1\pi$ bonds are present.



BOND ORDER

➢ It is defined as one half the distance between the no.of electrons present in bonding & anti bonding orbitals.

Bond order =
$$\frac{1}{2}[N_b - N_a]$$

> If B.O = +ve , then $N_b > N_a$ the molecule is stable.

> If B.O = -ve (or) 0, then $N_b < N_a$ the molecule is unstable.



NATURE OF BOND

The Bond order values of 1,2, 3 corresponds to a single, double (or) triple bonds respectively.



- ➢ If all the molecular orbitals in a molecule are doubly occupied (paired electrons), then the substance is diamagnetic.
- ➢ If one (or) more molecular orbitals are singly occupied (unpaired electron), it is paramagnetic.



- > The number of electrons in H_2 molecule is 2.
- > The electronic configuration is σ_{1S}^2

: B.O =
$$\frac{1}{2} [N_b - N_a] = \frac{1}{2} [2 - 0] = \frac{2}{2} = 1$$

 \rightarrow B.E = 438 K.J/mol⁻¹

$$\rightarrow$$
 B.L = 74 pm

- \rightarrow Diamagnetic
- > The two hydrogen atoms are bonded together by a single covalent bond.



ii) He₂:

- > The number of electrons in He_2 molecule is 4.
- > The electronic configuration is $\sigma_{1S}^2 \sigma_{1S}^{*2}^2$ $\Rightarrow B.O = \overline{2} [N_b - N_a] = \frac{1}{2} [2 - 2] = 0$ $\Rightarrow B.O = 0$
- > The molecule is unstable
- > The molecule does not exist
- > The bond order of He_2 is 0 Ans





iii) **B**₂:

- > The number of electrons in B_2 molecule is 10.
- > The electronic configuration is $\sigma_{1S} \sigma_{1S}^* \sigma_{2S} \sigma_{2S}^* \pi_{2Py} = \pi_{2Pz}$. 11 11 11 11 1

: B.O =
$$\frac{1}{2} [N_b - N_a] = \frac{1}{2} [6 - 4] = \frac{2}{2} = 1$$

The molecule contains two unpaired electrons, so it is paramagnetic in nature.



iv) C₂:

- > The number of electrons in C_2 molecule is 12.
- > The electronic configuration is $\sigma_{1S} \sigma_{1S}^* \sigma_{2S} \sigma_{2S}^* \pi_{2Py} = \pi_{2Pz}$. 11 11 11 11 11 11 11

: B.O =
$$\frac{1}{2} [N_b - N_a] = \frac{1}{2} [8 - 4] = \frac{4}{2} = 2$$

The molecule contains no unpaired electrons, so it is diamagnetic in nature.



v) N₂:

➤ The number of electrons in N₂ molecule is 14. $\sigma_{1s} < \sigma_{1s} < \sigma_{2s} < \sigma_{2s} < \sigma_{2s} < (\pi_{2Px} = \pi_{2Py}) < \sigma_{2Pz} < (\pi_{2Px} = \pi_{2Py}) < (\pi_{2Px} = \pi_{2Py}) < \sigma_{2Pz} <$

: B.O =
$$\frac{1}{2} [N_b - N_a] = \frac{1}{2} [10 - 4] = \frac{6}{2} = 3$$

> There is a triple bond between two nitrogen atoms.

> Due to the absence of unpaired electrons it is diamagnetic nature.

> In N₂ molecule $1\sigma \& 2\pi$ bonds are present.



Examples of molecular orbital

Explain the formation of O₂ on the basis of MOT

Ans : O₂ molecule is formed by the combination of two O atoms.

O = 8 electrons

 \therefore O₂ = 16 electrons





MOLECULAR ORBITAL DIAGRAM

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

Magnetic property = Paramagnetic

E.C. =KK
$$(\sigma 2s)^2(\sigma^* 2s)^2(\sigma 2p_z)^2$$

 $(\pi 2p_x^2) (\pi 2p_y^2) (\pi 2p_x^{-1}) (\pi^* 2p_y^{-1})$



\Rightarrow The stabilities of nitrogen molecule & it's ions are in the order.

$$N_2 > N_2^+ > N_2^- > N_2^{-2}$$



 \Rightarrow The stabilities of 'O₂' molecule & it's ion are

$$O_2^{+2} > O_2^+ > O_2^- > O_2^{-2}$$



Q) Write the electronic configuration and calculate the bond order of H_2^+ , H_2 & He_2^- . Explain why bond length in H_2^+ is longer than in H_2 .

Sol:

| Speci es | No.of e ⁻ | configuratio n | N b | Na | Bondorder = 1/2[N _b – N _a] |
|-----------------|-------------------------|---|--------|----|--|
| H_2^+ | 1 | σ_{1S}^{1} | 1 | 0 | 1/2 |
| H ₂ | 2 | σ_{1S}^{2} | 2 | 0 | 1 |
| He ₂ | 4 | $\begin{array}{c}\sigma_{1S}^{2}\\\sigma_{1S}^{*}^{2}\end{array}$ | 2 | 2 | 0 |



- The bond length in H₂⁺ is longer than H₂ because in H₂⁺ only one electron is present to shield.
- > The two nuclei form mutual repulsion.
- ➤ In H₂ there are 2e⁻ to hold the two nuclei thus nuclear repulsion is less than that in H₂⁺.

> Hence nuclear separation in H_2^+ is more than in H_2 .



Q) Write down the M.O diagram of O₂⁺, O₂, O₂^{-&} O₂⁻², Find out their B.O, magnetic property & stability.

Sol:

| Speci es | N.O of e ⁻ | B.O | Stability | | Magnetic nature |
|-----------------------|--------------------------|--|-----------|---|-----------------|
| O ₂ + | 15 | ¹ / ₂ [10- 5]=2.5 | eases | | Paramagnetic |
| O ₂ | 16 | ¹ / ₂ [10- 6]=2 | y decı | | Paramagnetic |
| O ₂ - | 17 | ¹ / ₂ [10- 7]=1.5 | abilit | | Paramagnetic |
| O ₂ -2 | 18 | ¹ /2[10- 8]=1 | S | , | Diaramagnetic |

Bond order: $O_2^+ > O_2 > O_2^- > O_2^{-2}$

Bond length: $O_2^+ < O_2 < O_2^- < O_2^{-2}$





1. Which is paramagnetic?





1. Which is not existed?



c) N₂
d) F₂



HYDROGEN BOND



HYDROGEN BOND:

The weak electrostatic force of attraction between partially positively charged hydrogen atom with highly electronegative (N, O, F) atom of same or different molecule is known as hydrogen bond.

> Hydrogen bond is weaker than covalent bond.



EX: HF: [Hydrogen fluoride]

In this molecule hydrogen bond exist between 'H' atom of one molecule and 'F' atom of another molecule.



➤ In the above example, hydrogen acts as bridge between 2 atoms which holds one atom by covalent bond and other by hydrogen bond.



- Hydrogen bond is represented by dotted line(-----) while a solid line represents the covalent bond.
- The physical state of the compound depends on the magnitude of 'Hbonding'. It is maximum in the solid state and minimum in gaseous state.
- ➤ The hydrogen bonds have strong influence on the structure & properties of the compounds.

TYPES OF HYDROGEN BOND:

- > Inter molecular hydrogen bond
- > Intra molecular hydrogen bond



INTER MOLECULAR HYDROGEN BOND

It is formed between two different molecules of the same or different compounds.





INTRA MOLECULAR HYDROGEN BOND

> This type of hydrogen bond is formed with in the molecule .



O – nitro phenol

Salicylaldehyde



Important points:

- In the formation of hydrogen bond, there are dipole-dipole attraction at opposite ends.
- > Hydrogen bond formation never involves more than '2' atoms.
- Strength of hydrogen bond increases with difference in Electro negativities of elements H - F>H - O>H - N
- > All the '3' atoms in X H-----X hydrogen bond, lies in one plane.
- H- bond may exist in solid phase, Liquid phase & in solution, also may exist in gas phase.



Important points:

- > Strength of the hydrogen bond in solids, liquids, gases
- Solids > Liquids > Gases
- > Hydrogen bonding can be detected by using following techniques.
 - a) Dipole moment.
 - b) Solubility behaviour.
 - c) Elevation in B.pt & Depression in F.pt.
 - d) Enthalpy of mixing.
 - e) IR spectrum.
- Weak hydrogen bond has been reported in following system also.



EX:

- Cl has the same electronegative as nitrogen but does not form strong hydrogen bonds.
- because the Cl atom is much larger than N-atom & so it's electrostatic attractions are weak.
- Deuterium also forms H-bond, which may be stronger (or) weaker than 'H'.
- Stronger hydrogen bond exists in KHF₂ between H and F H -----F⁻ (B.E ≅ 113 K.J/mol).



Explanations based on H-bonding:

- Intermolecular hydrogen bonding increases the B.P and makes a substance less volatile, the effect of intramolecular hydrogen bonding is reverse.
- > B.P of $H_2O > B.pt$ of H_2S , i.e., H_2O is a liquid, while H_2S is a gas.
- ▶ **B.P of NH**₃ > **B.pt of PH**₃.
- ➢ B.P of HF > B.pt of HCl.
- > B.P of amines are in the order $1^0 > 2^0 > 3^0$.



Explanations based on H-bonding:

Glycerol is viscous in nature and has very high B.Pt (290°C), because one molecule of glycerol forms 3- inter molecular H-bonds with another glycerol molecule.





1. Strong H-bond is present in...

a) solidsb) liquids

c) gases

d) none



2. More boiling point is present in...

a) 1⁰- amines
b) 2⁰- amines

- c) 3⁰- amines
- d) all are same



SOLUBILITY



Solubility:

➢ Intermolecular hydrogen bonding increases the solubility of a covalent compound in water which can from H-bond with it.

Ex: Solubility of alcohols in H₂O

> Intramolecular hydrogen bonding decreases the solubility.

Ex: O-nitrophenol and salicylic acid are less soluble in H₂O than corresponding Para-isomers.



High viscosity:

Inter molecular hydrogen bonding makes a compound more viscous and that is why glycerol is highly viscous liquid.

H-bonding in biological systems also play an important role and stability of proteins and nucleic acids is due to hydrogen bonding.





1. Which one would have higher entropy, if same amounts of two are taken.

a) O-nitrophenol
b) P-nitrophenol
c) Both (a) and (b)
d) Can't be compared

> Due to intra molecular hydrogen bonding there will be more entropy.


- 2. Ammonium salts are more soluble in H_2O than the corresponding sodium salts explain:
- H-atom in NH₄⁺ shows hydrogen bonding with water molecule in addition to dipole-ion attraction where as Na⁺ does not.



3. H₂O is a liquid, while H₂S is a gas at ordinary temperature. Explain?

> Due to hydrogen bonding between O and H atoms in H_2O molecule.



4. Alcohols have higher boiling point than corresponding alkane and thiol explain:

> H-bonding in alcohols



- 5. Glycerol is more viscous than ethane explain?
 - > Due to extensive H-bonding in glycerol.



- 6. Thio-ethers have higher boiling point than ethers explain:
- The molecular mass of thio-ethers (R-S-R) are higher than ethers(R-O-R) and thus thio-ethers have high B.P



7. CH₄ and H₂O have nearly same molecular mass. Yet CH₄ has a B.Pt 112K and water 373k explain:

➢ H₂O shows H-bonding which results in higher (abnormal) B.Pt



8. A compound with two (or) more – 'OH' groups present on a carbon is usually unstable, but chloral hydrate CCl₃CH (OH)₂ is fairly stable.

> Chloral hydrate shows intra molecular H- bonding.





- 9. Give an example where H-bond exists in gas phase
- ➢ In vapour phase acetic acid exists as dimer due to H-bonding. Hence acetic acid molecular weight is 120 in vapour phase.





10. Explain the H-bond formed b/w NH₃ and H₂O





11. Water has the highest boiling point among NH₃, HCl and H₂O why?

> Due to strongest H-bonding in H₂O



12. Which of the following does not forms H – bonding ?

a) HF
b) H₂O
c) NH₃
d) HCl



13. Which is an example of intramolecular H-bonding?

(a) HF

(b) Water

(c) Salicylic acid(d) Ethyl alcohol



14. H – bonding is formed between _____.

(a) atoms
(b) ions
(c) molecules
(d) metals