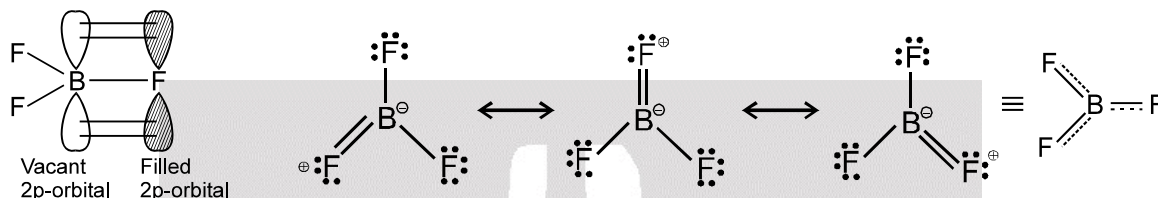




# Chemical Bonding-III

## Section (A) : Back bonding

Back bonding generally takes place when out of two bonded atoms one of the atom has vacant orbitals (generally this atom is from second or third period) and the other bonded atom is having some non-bonded electron pair (generally this atom is from the second period). Back bonding increases the bond strength and decreases the bond length. For example, in  $\text{BF}_3$  the boron atom completes its octet by accepting two 2p-electrons of fluorine into 2p empty orbital.

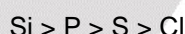


- Decrease in B-F bond length is due to delocalised  $p\pi-p\pi$  bonding between filled p-orbital of F atom and vacant p-orbital of B atom.

The extent of back bonding is much larger if the orbitals involved in the back bonding are of same size, for example the extent of back bonding in boron trihalides is as follows :  $\text{BF}_3 > \text{BCl}_3 > \text{BBr}_3$

There is  $p\pi-p\pi$  back bonding in boron trihalide. The extent of back bonding decreases from  $\text{BF}_3$  to  $\text{BI}_3$  because of increasing size of p-orbitals participating in back bonding that is from 2p (in F) to 4p (in Br).

- The extent of back bonding decreases if the atom having vacant orbitals is also having some non-bonded electron pairs on it. So among the atoms of third period the extent of back bonding follows the order :

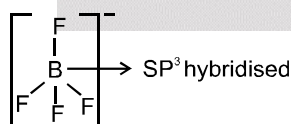


- The extent of  $p\pi-p\pi$  overlapping  $\propto \frac{1}{\text{Lewis acid character}}$

## Solved Examples

**Ex-1.** Compare B-F bond length in  $\text{BF}_3$  and  $[\text{BF}_4]^-$ .

**Sol.**  $\text{BF}_3 + \text{KF} \longrightarrow \text{K}^+ [\text{BF}_4]^-$

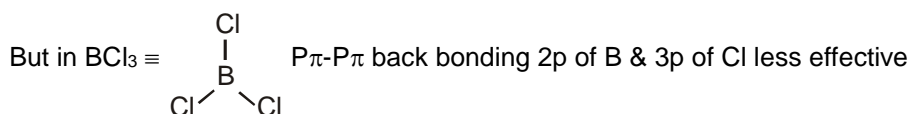


[No back bonding]

B-F Bond length (order)  $[\text{BF}_4]^- > \text{BF}_3$

**Ex-2.** Explain all boron trihalides are Lewis acids also explain their order.

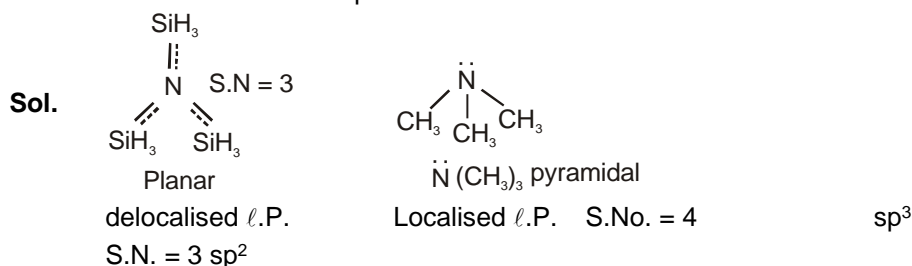
**Sol.** Boron trihalides are electron deficient molecules therefore act as a Lewis acids



tendency to accept L.P in  $\text{BCl}_3 > \text{BF}_3$ . Lewis acid strength  $\text{BF}_3 < \text{BCl}_3 < \text{BBr}_3$ .

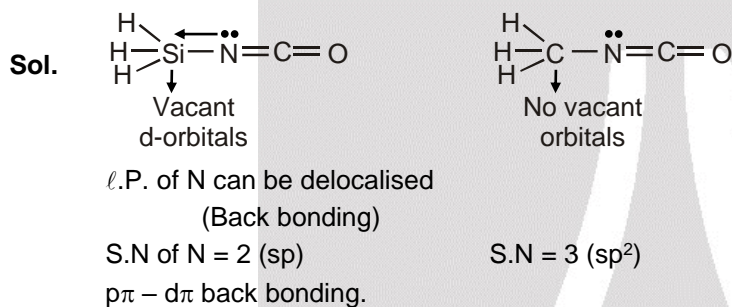


**Ex-3.** Trisilylamine is a planar molecular and does not act as a lewis base while trimethyl is a pyramidal and act as lewis base. Explain it :



\* But in a similar compound  $\text{N}(\text{PH}_3)_3$  the shape is found to be pyramidal, so N atom must be  $sp^3$  hybridised due to much less extent of back bonding into the vacant orbitals of P.

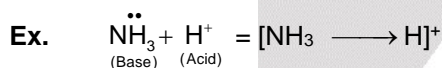
**Ex-4.** Silyl isocyanate ( $\text{SiH}_3\text{NCO}$ ) is linear but methyl isocyanate ( $\text{CH}_3\text{NCO}$ ) is bent explain !



## Section (B) : Electron deficient bonding

### Definition of Lewis acid and base :

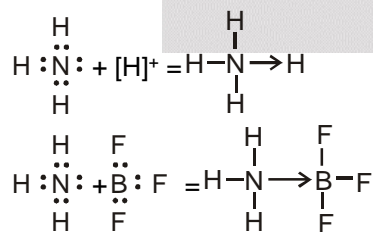
- A lewis acid is an electron pair acceptor and lewis base is electron pair donor.



Proton is a lewis acid and Ammonia is lewis base since the lone pair of electron of the nitrogen atom can be donated to a proton.

### Coordinate Bond :

A bond in which sharing of a pair of electron between two atoms, where both electrons originate from one atom and none from other is called **coordinate bond**.



### Electron deficient bonding :

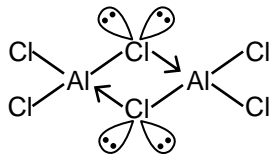
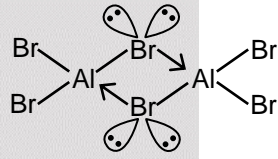
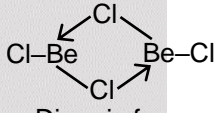
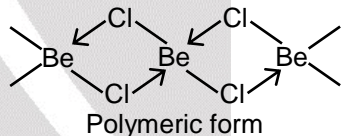
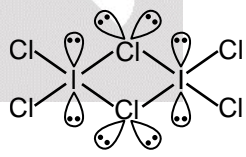
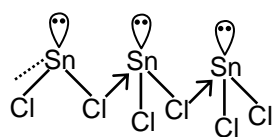
There are many compounds in which some electron deficient bonds are present apart from normal covalent bonds or coordinate bonds which are 2c-2e bonds (two centre two electron bonds). These electron deficient bonds have less number of electrons than the expected such as three centre-two electron bonds(3c-2e) present in diborane  $\text{B}_2\text{H}_6$ ,  $\text{Al}_2(\text{CH}_3)_6$ ,  $\text{BeH}_2(\text{s})$ , bridging metal carbonyls.

$\text{B}_2\text{H}_6$  (diborane)  $\Rightarrow$  It is having (3centre – 2electron) bond / (banana bond) / (electron deficient bond)



Compound	Bridge	Hybridisation	Planar/Non Planar	Structure
$B_2H_6$	$3c - 2e^-$	$sp^3-s-sp^3$	Non planar	
	<p><math>B_2H_6</math> have 4 <math>2c-2e</math> bonds and 2 <math>3c-2e</math> bonds. Bridging bonds have larger bond length than terminal bonds. Angle between terminal bonds is more than angle between bridging bonds if all 4 terminal bonds are in one plane then bridging bonds are in perpendicular plane.</p>			
$Al_2H_6$	$3c - 2e^-$	$sp^3-s-sp^3$	Non planar	
$Be_2H_4$	$3c - 2e^-$	$sp^2-s-sp^2$	planar	
$(BeH_2)_n$	$3c - 2e^-$	$sp^3-s-sp^3$	Non planar	
	All bonds except the extreme ends are $3C - 2e^-$ bonds.			
$Al_2(CH_3)_6$	$3c - 2e^-$	$sp^3-sp^3-sp^3$	Non planar	
$Al_2(C_6H_5)_6$	$3c - 2e^-$	$sp^3-sp^2-sp^3$	Non planar	



Compound	Bridge	Hybridisation	Planar/ Non Planar	Structure
$\text{Al}_2\text{Cl}_6$	$3c - 4e^-$	$sp^3-sp^3-sp^3$	Non Planar	 <p>(Driving force is to complete the octet)</p> <p>Only covalent bond, no electron deficient bond is present.</p>
$\text{Al}_2\text{Br}_6$	$3c - 4e^-$	$sp^3-sp^3-sp^3$	Non Planar	 <p>Dimeric form</p>
$\text{Be}_2\text{Cl}_4$	$3c - 4e^-$	$sp^2-sp^3-sp^2$	Planar	 <p>Dimeric form</p>
$(\text{BeCl}_2)_n$ (solid)	$3c - 4e^-$	$sp^3-sp^3-sp^3$	Non Planar	 <p>Polymeric form</p>
$\text{I}_2\text{Cl}_6$	$3c - 4e^-$	$sp^3d^2-sp^3-sp^3d^2$	Planar	 <p>Dimeric form</p>
$\text{SnCl}_2(\text{Solid})$	$3c - 4e^-$	$sp^3-sp^3-sp^3$	Non Planar	 <p>(Driving force is to complete the octet)</p> <p>Polymeric form</p>



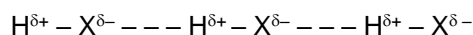


## Section (C) : H-bonding & intermolecular force of attraction.

### Hydrogen Bond :

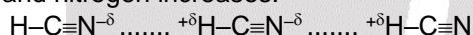
Discovered by Huggius, Latimer and Rodbush.

When hydrogen is bonded to strongly electronegative element 'X'. The electron pair shared between the two atoms moves far away from hydrogen atom. As a result the hydrogen atom becomes highly electropositive with respect to the other atom 'X'. Since there is displacement of electrons towards X, the hydrogen acquires fractional positive charge ( $\delta^+$ ) while 'X' attain fractional negative charge ( $\delta^-$ ). This results in the formation of a polar molecule having electrostatic force of attraction which can be represented as :

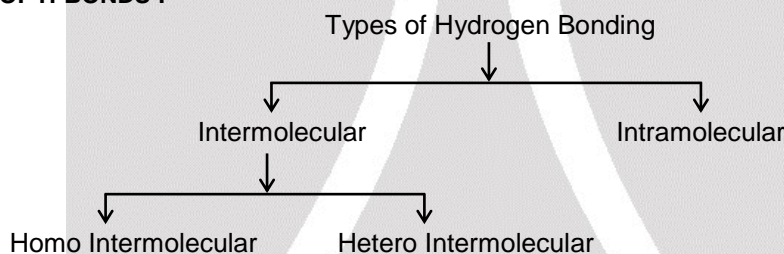


### Main condition for Hydrogen bonding :

- Hydrogen should be covalently bonded with high electronegative element like F, O, N.
- Atomic size of electronegative element should be small  
Decreasing order of atomic size is :  $N > O > F$   
Decreasing order of electronegativity is :  $F (4.0) > O (3.5) > N (3.0)$
- Strength of Hydrogen bond  $\propto$  Electronegativity of element  $\propto \frac{1}{\text{atomic size of element}}$
- Hydrogen bonding occurs in HCN, due to ( $-C \equiv N$ ) triple bond (sp hybridisation), electronegativities of carbon and nitrogen increases.



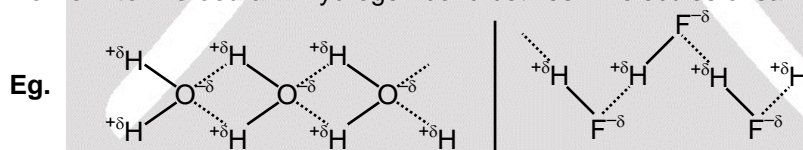
### TYPES OF H-BONDS :



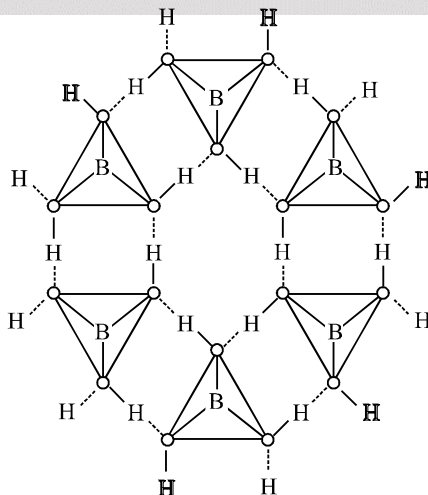
### (A) Intermolecular Hydrogen bond :

Hydrogen bond formation between two or more molecules of either the same or different compounds known as **Inter molecular Hydrogen bonding**.

- Homo intermolecular** : Hydrogen bond between molecules of same compounds.



Eg. Boric acid

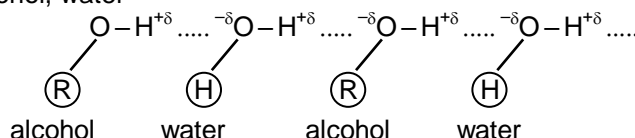




In the solid state, the  $B(OH)_3$  units are hydrogen bonded together in to two dimensional sheets with almost hexagonal symmetry. The layered are quite a large distance apart ( $3.18 \text{ \AA}$ ).

(ii) **Hetro intermolecular** : Hydrogen bond between molecules of different compounds.

Eg. alcohol, water



(B) **Intra molecular Hydrogen bond** : It takes place within the molecule.

(i) Hydrogen bonded with electronegative elements of a functional group, form Hydrogen bond with another electronegative element present on nearest position on the same molecule.

(ii) This type of Hydrogen bond is mostly occured in organic compounds.

(iii) It result in ring formation (Chelation).

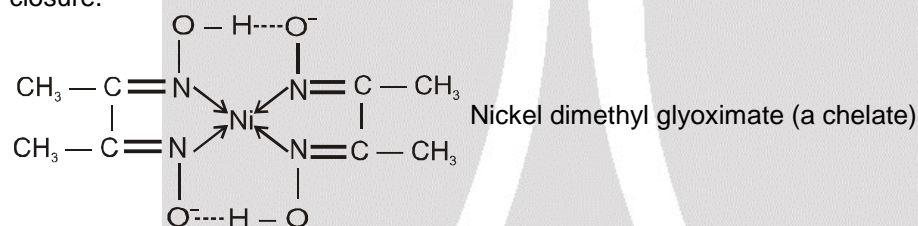
**Necessary conditions for the formation of intramolecular hydrogen-bonding :**

(a) the ring formed as a result of hydrogen bonding should be planar.

(b) a 5- or 6-membered ring should be formed.

(c) interacting atoms should be placed in such a way that there is minimum strain during the ring closure.

Ex-1.



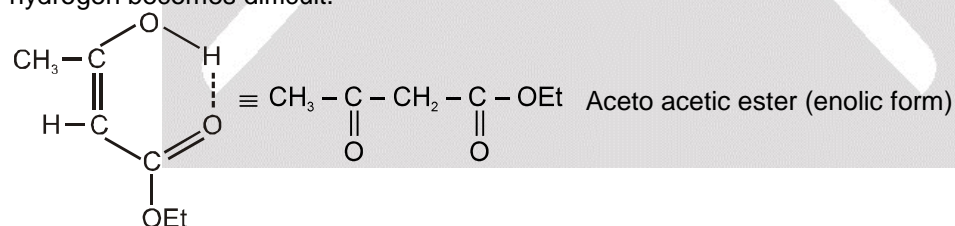
Extra stability of the complex is because of intramolecular hydrogen bonding in addition to the chelating effect.

Ex-2.



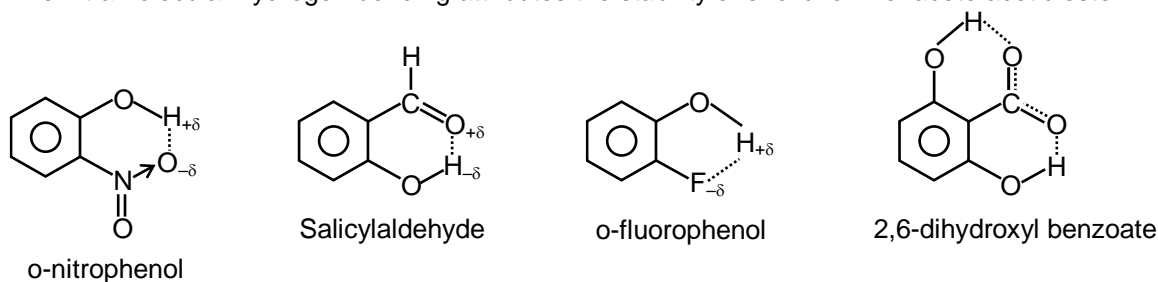
$K_1$  of peroxomono sulphuric acid (i.e., caros acid) is greater than  $K_2$ . After the loss of one hydrogen, the persulphate ion gets stabilised due to intramolecular hydrogen bonding and thus the removal of second hydrogen becomes difficult.

Ex-3.



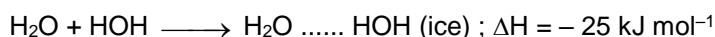
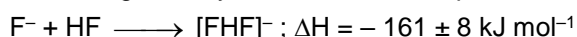
The intramolecular hydrogen bonding attributes the stability of enolic form of aceto acetic ester.

Ex-4.



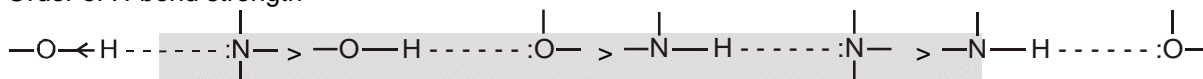
**Strength of H-bond :**

The strength of H-bond is usually very low (5-10 kJ/mol) but in some cases this value may be as high as 50 kJ/mol. The strongest H-bonds are formed by F atoms. Deuterium is more electropositive than H, therefore it also form stronger bonds. The strength of the H-bond can be compared by the relative bond energies and the geometry of the various compounds as given below.



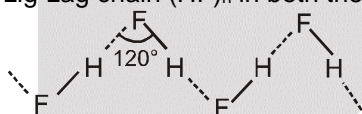
The magnitude of H-bonding depends on the physical state of the compounds. H-bonding is maximum in the solid state and minimum in the gaseous state. Thus hydrogen bonds have strong influence on the structure and properties of the compounds.

○ Order of H-bond strength



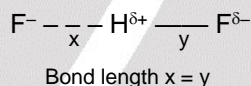
Hydrogen bond is weaker than covalent bond.

The hydrogen bonds in HF link the F atom of one molecule with the H-atom of another molecule, thus forming a zig-zag chain  $(\text{HF})_n$  in both the solid and also in the liquid.

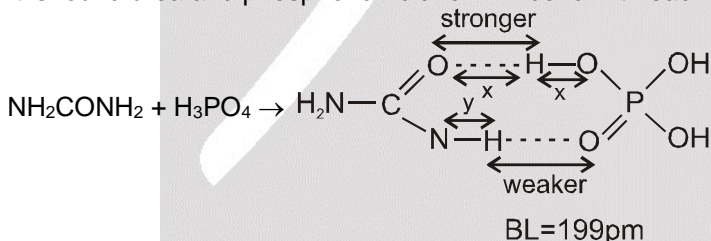


Some hydrogen bonding also occurs in the gas, which consists of a mixture of cyclic  $(\text{HF})_6$  polymers, dimeric  $(\text{HF})_2$ , and monomeric HF.

Very strong hydrogen bonding occurs in the alkali metal hydrogen fluorides of formula  $\text{M}[\text{HF}_2]$ ; in  $\text{KHF}_2$ , for example, an X-ray diffraction study together with a neutrons diffraction study shows that there is a liner symmetrical anion having an over all, F-H-F distance of 2.26 Å, which may be compared with the H-F bond length of 0.92Å in hydrogen fluoride monomer.



It is found urea and phosphoric Acid form H bond with each other



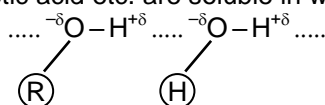
2 types of H-bond

B.L.OH  $\Rightarrow x = 122 \text{ pm}$  ;

B.L. N-H  $\Rightarrow y = 100 \text{ pm}$

**Effect of Hydrogen bond on physical properties :**(i) **Solubility :****(A) Inter molecular Hydrogen bonding**

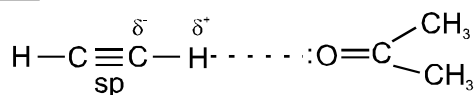
(a) Few organic compounds (Non-polar) are soluble in water (Polar solvent) due to Hydrogen bonding. e.g alcohol, acetic acid etc. are soluble in water.



Other examples- Glucose, Fructose etc, dissolve in water.

(b) Ketone, ether, alkane etc. are insoluble (no Hydrogen bond)

But  $\text{C}_2\text{H}_2$  is highly soluble in acetone due to H bonding



None of the above 2 has H-bonding individually.

But  $\text{C}_2\text{H}_2$  is not soluble in water because water molecules already so much associated through H-bond that it is almost impossible for  $\text{C}_2\text{H}_2$  molecules to break that association.

So, it is not soluble in  $\text{H}_2\text{O}$  (l)

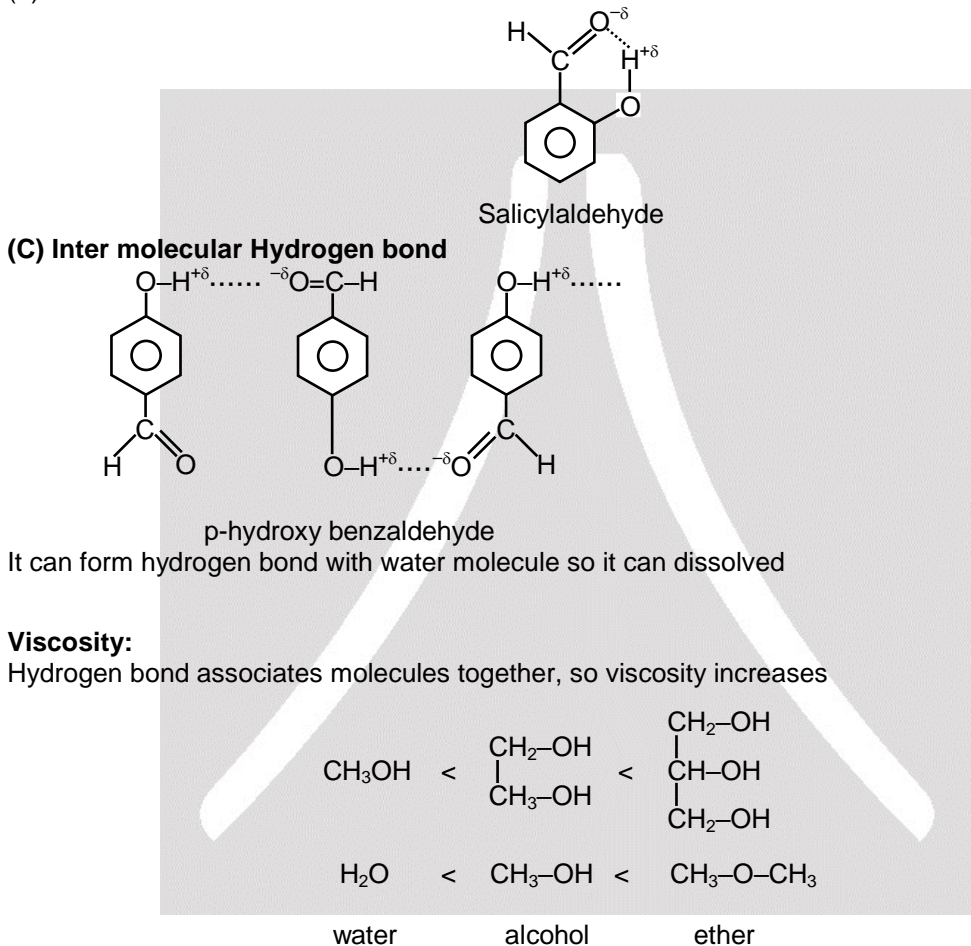
(c) Solubility order-  $\text{CH}_3\text{OCH}_3 < \text{CH}_3\text{OH}$

Primary amine > secondary amine > tertiary amine

**(B) Intra molecular Hydrogen bonding :**

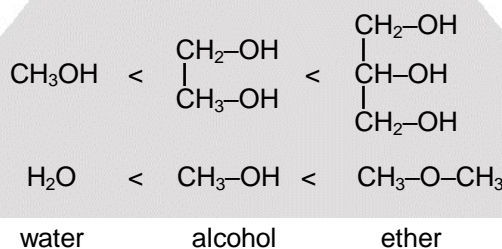
(a) It decreases solubility as it form chelate by Hydrogen bonding, so hydrogen is not free for other molecule.

(b) It can't form H-bond with water molecule so can't dissolves.



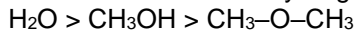
**(ii) Viscosity:**

Hydrogen bond associates molecules together, so viscosity increases



**(iii) Melting point and boiling point**

(a) Due to intermolecular hydrogen bond melting point & boiling point of compounds increases.

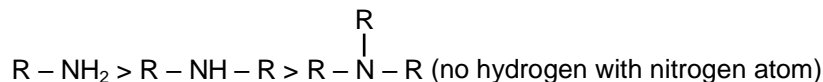


(b) Trihydric alcohol > dihydric alcohol > monohydric alcohol

Monocarboxylic acid form stronger hydrogen bond than alcohol of comparable molecular weight. Therefore boiling point of carboxylic acid is higher than alcohol.

(c) Decreasing order of melting point & boiling point isomer amines-

1°-amine > 2°-amine > 3°-amine





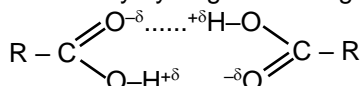


- (d) Boiling points of VA, VIA, VIIA hydrides decreases on decreasing molecular weights.
- | VA               | VIA              | VIIA | boiling point   |
|------------------|------------------|------|---|
| NH <sub>3</sub>  | H <sub>2</sub> O | HF   | HF > HI > HBr > HCl   |
| PH <sub>3</sub>  | H <sub>2</sub> S | HCl  | H <sub>2</sub> O > TeH <sub>2</sub> > SeH <sub>2</sub> > H <sub>2</sub> S |
| AsH <sub>3</sub> | SeH <sub>2</sub> | HBr  | NH <sub>3</sub> > SbH <sub>3</sub> > AsH <sub>3</sub> > PH <sub>3</sub>   |
| SbH <sub>3</sub> | TeH <sub>2</sub> | HI   |   |

But sudden increase in boiling point of NH<sub>3</sub>, H<sub>2</sub>O and HF is due to hydrogen bonding  
 H<sub>2</sub>O > HF > NH<sub>3</sub>

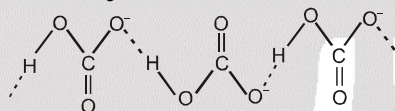
- (e) Intramolecular hydrogen bonding gives rise to ring formation, so the force of attraction among these molecules are vander waal's force. So, melting point and boiling point are low.

- (iv) **Molecular weight:** Molecular weight of CH<sub>3</sub>COOH is double of its molecular formula, due to dimer formation occur by hydrogen bonding.

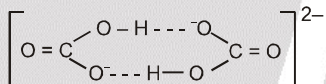


The crystal structures of NaHCO<sub>3</sub> and KHCO<sub>3</sub> both show hydrogen bonding, but are different.

- (a) In NaHCO<sub>3</sub>, the HCO<sub>3</sub><sup>-</sup> ions are linked into an infinite chain.



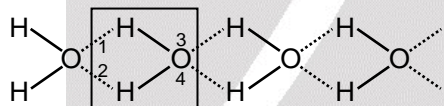
- (b) in KHCO<sub>3</sub>, RbHCO<sub>3</sub>, CsHCO<sub>3</sub>, HCO<sub>3</sub><sup>-</sup> forms a dimeric anion.  
 Solubility in water NaHCO<sub>3</sub> < KHCO<sub>3</sub> < RbHCO<sub>3</sub> < CsHCO<sub>3</sub>



- (v) **Physical state:** H<sub>2</sub>O is liquid while H<sub>2</sub>S is gas.

- (a) **Water and Ice :** Both have hydrogen bonding even then density of ice is less than water. Volume of ice is more because of open cage like crystal structure, from by association of water molecules with the help of hydrogen bond.

H<sub>2</sub>O becomes solid (Ice) due to four hydrogen bond among water molecule are formed in tetrahedral manner.

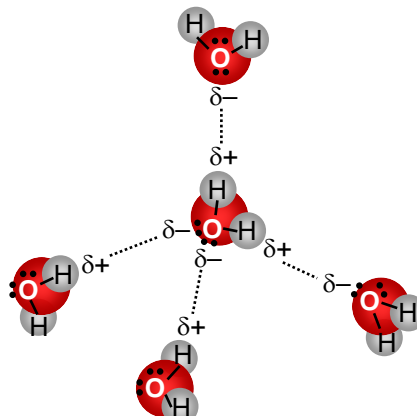
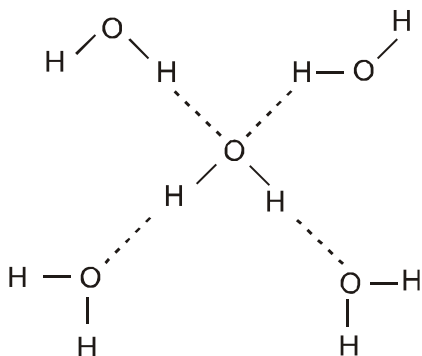


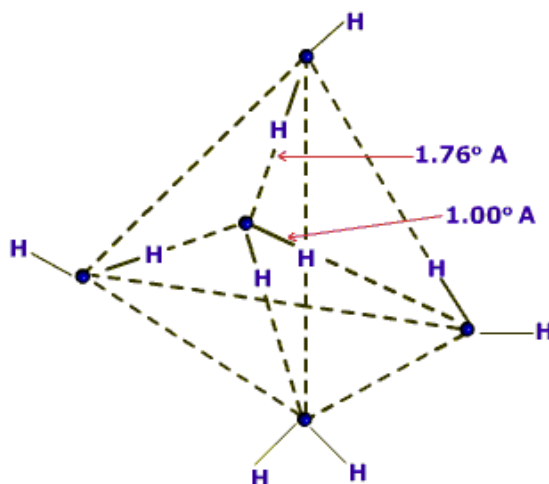
- (b)
- |   |   |
|---|---|
| $\begin{array}{c} \text{H}-\text{O} \cdots \text{H}-\text{O} \\   \quad \quad   \\ \text{H} \quad \quad \text{H} \end{array}$ | $\begin{array}{c} \text{D}-\text{O} \cdots \text{D}-\text{O} \\   \quad \quad   \\ \text{D} \quad \quad \text{D} \end{array}$ |
| $E_{\text{H}^+/\text{H}_2}^\circ = 0$   | $E_{\text{D}^+/\text{D}_2}^\circ = -0.01\text{V}$   |

D is more electro positive than H

So, it is capable of forming slightly more H bond than H that why D<sub>2</sub>O molecule are more associated and its boiling point is more than the boiling point of H<sub>2</sub>O liquid.

- (c) **Ice floats on water**





No. of water molecule attached to 1  $\text{H}_2\text{O}$  molecule = 4

No. of H bonds in a molecule = 2

One  $\text{H}_2\text{O}$  is tetrahedrally bonded with 4  $\text{H}_2\text{O}$  molecules.

Intermolecular spacing increases

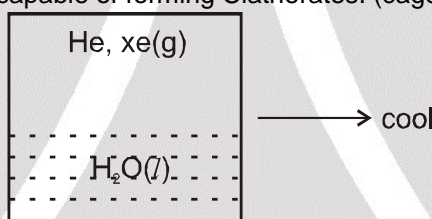
So, density of ice decreases

**(d)  $\text{D}_2\text{O}(\text{s})$**

Density of  $\text{D}_2\text{O}(\text{s}) >$  density of  $\text{H}_2\text{O}(\text{l})$

So, it sinks in  $\text{H}_2\text{O}(\text{l})$  but floats on  $\text{D}_2\text{O}(\text{l})$

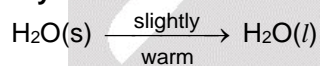
Due to open structure, ice is capable of forming Clathrates. (cage like compounds)



Xe will be captured by ice.

Xe(g) because of its bigger atomic size is trapped in the cages formed by  $\text{H}_2\text{O}$  molecules in the structures of ice. He cannot be trapped due to smaller size. Such compounds are called clathrates compounds.

**(e) Density of water max at  $4^\circ\text{C}$**

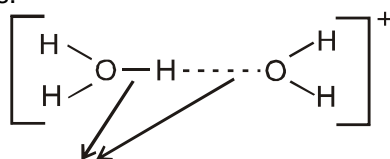


ice  $T < 4^\circ\text{C}$   
density decreases  $> 0^\circ\text{C}$

In this temperature region some of the ice melts and hence some  $\text{H}_2\text{O}$  molecule go into the cages of remaining ice structure.

So, volume decreases density increases becoming max at  $4^\circ\text{C}$  but beyond this temp thermal effects become dominating volume increases then density decreases.

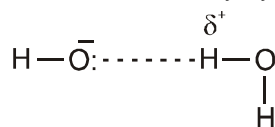
**(f)** An interesting hydrate of Hydronium ion with the formula  $\text{H}_3\text{O}^+(\text{H}_2\text{O})_{20}$  in which  $\text{H}_2\text{O}$  molecules are hydrogen bonded. Some lower species like  $\text{H}_5\text{O}_2^+$  have also been observed where the 2 water molecules are linked three H bonds.



Bond strength 100-150 kJ/mole  
(Proton transfer takes place)

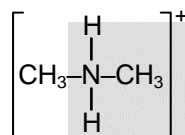
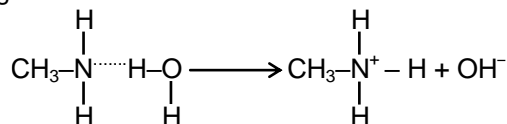


Similarly  $\text{H}_3\text{O}_2^-$  ions are also observed which is actually hydrate  $\text{OH}^-$  ion.



**(vi) Base strength**

$\text{CH}_3\text{NH}_2$ ,  $(\text{CH}_3)_2\text{NH}$ ,  $(\text{CH}_3)_3\text{N}$ , form hydrogen bond with water. So, less hydrolysis i.e. it gives  $\text{OH}^-$  ions. While  $(\text{CH}_3)_4\text{N}^+$   $\text{OH}^-$  (ammonium compound) will give  $\text{OH}^-$  ion in large amount due to no hydrogen bonding.



(Ammonium compound)  
No Hydrogen atom bonded directly with Nitrogen atom so no hydrogen bonding occurs.

**Ex.** Which is a stronger base why?

(a) Trimethyl ammonium hydroxide  $[\text{N}(\text{CH}_3)_3]\text{OH}$

(b) Tetramethyl ammonium hydroxide  $[\text{N}(\text{CH}_3)_4]^+\text{OH}^-$

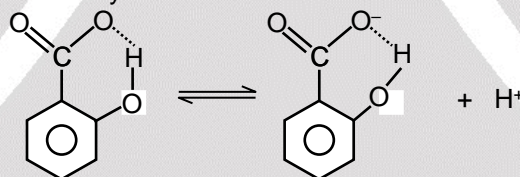
b is more basic  
 $(\text{CH}_3)_3\text{N} : \cdots \cdots \text{H}-\text{O}$

Due to intramolecular H bonding, release of  $\text{OH}^-$  ion become difficult  $\therefore$  it is a weaker base.

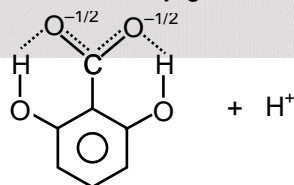
**Effect of intramolecular H-bonding**

**(i) Strength of acid :**

(a) The formation of intramolecular H-bonding in the conjugate base of an acid gives extra stability to conjugate base and hence acid strength increases eg. Salicylic acid is stronger than benzoic acid and 2, 6-dihydroxy benzoic acid > salicylic acid.



Conjugate base



2,6-dihydroxyl benzoate

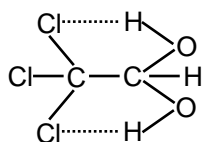
(b)  $\text{C}_2\text{H}_5\text{SH}$  is more acidic than  $\text{C}_2\text{H}_5\text{OH}$ . In  $\text{C}_2\text{H}_5\text{OH}$ , hydrogen bonds forms, so  $\text{H}^+$  is not free

(c)  $\text{HF}$  is weaker acid than  $\text{HI}$ , due to hydrogen bond in  $\text{H}-\text{F}$ ,  $\text{H}^+$  is not free.

**(ii) Stability of chloral hydrate :**

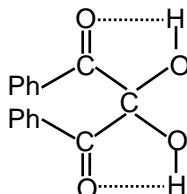
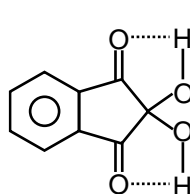
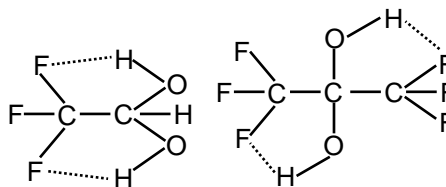
If two or more  $\text{OH}$  group on the same atom are present it will be unstable, but chloral hydrate is stable (due to Hydrogen bonding)



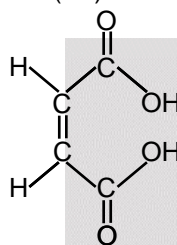


Chloral hydrate

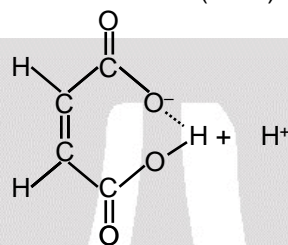
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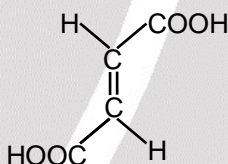
(iii) Maleic acid (cis) is stronger acid than fumaric acid (trans).



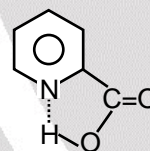
(Maleic acid)



Stable conjugate base of maleic acid



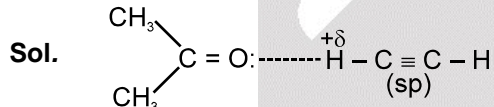
Fumaric (No-intramolecular hydrogen bonding)



**Note :** The relative strength of various bonds is as follows  
 Ionic bond > covalent bond > Metallic bond > Hydrogen bond > Vander waal's bond.

### Solved Examples

**Ex-5.**  $C_2H_2$  is not soluble in  $H_2O$  but it is highly soluble in acetone.



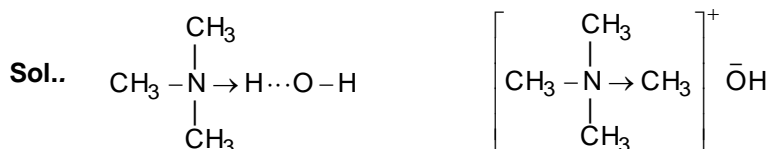
In hybridisation as %s character increase, electronegativity increase hence  $C_2H_2$  forms H-bonds with O-atom of acetone and get dissolved. But  $H_2O$  molecules are so much associated that it is not possible for  $C_2H_2$  molecules to break that association, hence  $C_2H_2$  is not soluble in  $H_2O$ .

**Ex-6.** Why crystalline sodium peroxide is highly hygroscopic in nature.

**Sol.**  $Na_2O_2$  forms stable hydrates on account of H-bonding.



**Ex-7.** Explain that tetramethyl ammonium hydroxide is a stronger base than that of trimethyl ammonium hydroxide.







In the trimethyl compound the O–H group is hydrogen bonded to Me<sub>3</sub>NH group and this makes it more difficult for the OH group to ionize and hence it is a weak base.

In the tetramethyl compound, hydrogen bonding cannot occur, so the OH<sup>-</sup> group ionizes easily and thus it is a much stronger base.

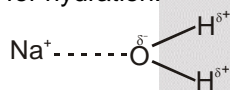
### Intermolecular forces (Van der Waal's Forces) :

Intermolecular attractions hold two or more molecules together. These are weakest chemical forces and can be of following types.

- (a) Ion-dipole attraction (b) Dipole-dipole attraction (Keesom forces)  
 (c) Ion-induced dipole attraction (Debye force) (d) Dipole-induced dipole attraction  
 (e) Instantaneous dipole- Instantaneous induced dipole attraction (Dispersion force or London forces)

**Strength of van der Waal's forces**  $a > b > c > d > e$

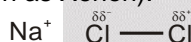
- (a) **Ion-dipole attraction** : Exists between an ion and a polar molecule. Its strength depends on (i) size of ion (ii) charge on the ion (iii) dipole moment of the polar molecule. It is thought to be directional. Ion-dipole forces are important in solutions of ionic compounds in polar solvents where solvated species such as Na(OH<sub>2</sub>)<sub>x</sub><sup>+</sup> and F(H<sub>2</sub>O)<sub>y</sub><sup>-</sup> (for solution of NaF in H<sub>2</sub>O) are found. Hence this force is responsible for hydration.



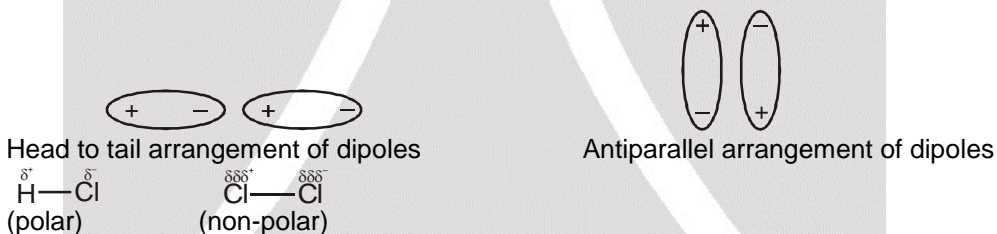
- (b) **Dipole-dipole attraction** : This is electrostatic attractions between the oppositely charged ends of permanent dipoles. Exists between polar molecules and due to this force gas can be liquefied.



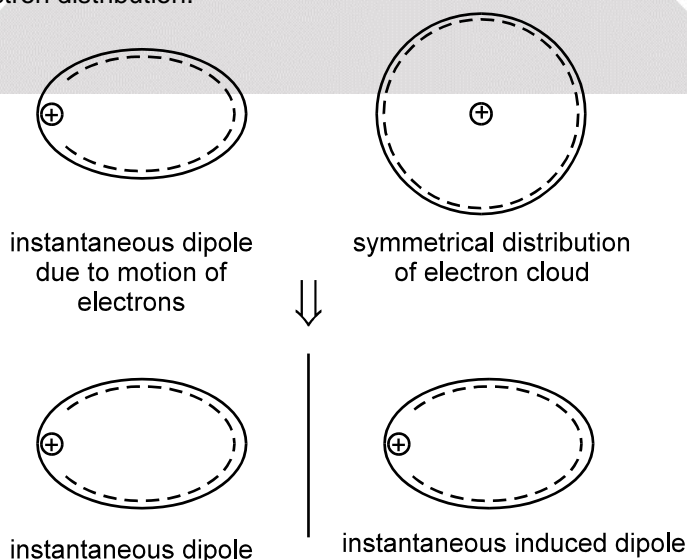
- (c) **Ion-induced dipole attraction** : Exists between ion and non-polar molecules (e.g., an atom of a noble gas such as Xenon).



- (d) **Dipole-induced dipole attraction** : Exists between polar and non-polar molecules.



- (e) **Instantaneous dipole- Instantaneous induced dipole attraction** : Exists among the non-polar molecules like H<sub>2</sub>, O<sub>2</sub>, Cl<sub>2</sub> etc. in solid or liquid states. Even in atoms in molecules which have no permanent dipole, instantaneous dipoles will arise as a result of momentary unbalances in electron distribution.





London forces are extremely short range in action and the weakest of all attractive forces. The London forces increase rapidly with molecular weight, or more properly, with the molecular volume and the number of polarizable electrons.

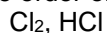


**Note :** Fluoro carbon have usually low boiling points because tightly held electrons in the fluorine atoms have a small polarizability.

- Strength of vander waal force  $\propto$  molecular mass.
- van der Waal's force  $\propto$  boiling point.

### Solved Examples

**Ex-8.** Give the order of boiling point of following



**Sol.**  $\text{Cl}_2-\text{Cl}_2$  <  $\text{HCl}-\text{HCl}$  (boiling point)  
 dispersion force dipole-dipole attraction  
 As dipole-dipole attraction is stronger than dispersion force.

**Ex-9** Arrange the inert gases, according to their increasing order of boiling points

**Sol.**  $\text{He} < \text{Ne} < \text{Ar} < \text{Kr} < \text{Xe}$  (boiling point)

Because strength of van der Waal's force increases down the group with increase in molecular mass.