BONDING

LECTURE 4

INTRODUCTION

- Compounds are formed when two or more elements combine.
- In any compound, the constituent elements are linked to each other by chemical bonds.
- There are three types of chemical bonds; namely **ionic or electrovalent bond covalent bond, and Coordinate bond or dative covalent bond.**

Ionic Bond

- This type of bond formation is due to the complete transfer of electrons between atoms that form compounds.

- Let us consider the case where NaCl is formed from the two elements, Na and Cl.

- Sodium loses an electron to form Na^+ cation while Chlorine gains an electron to form Cl^- anion

There has been the transfer of electron between Na and Cl to form oppositely charged ions.

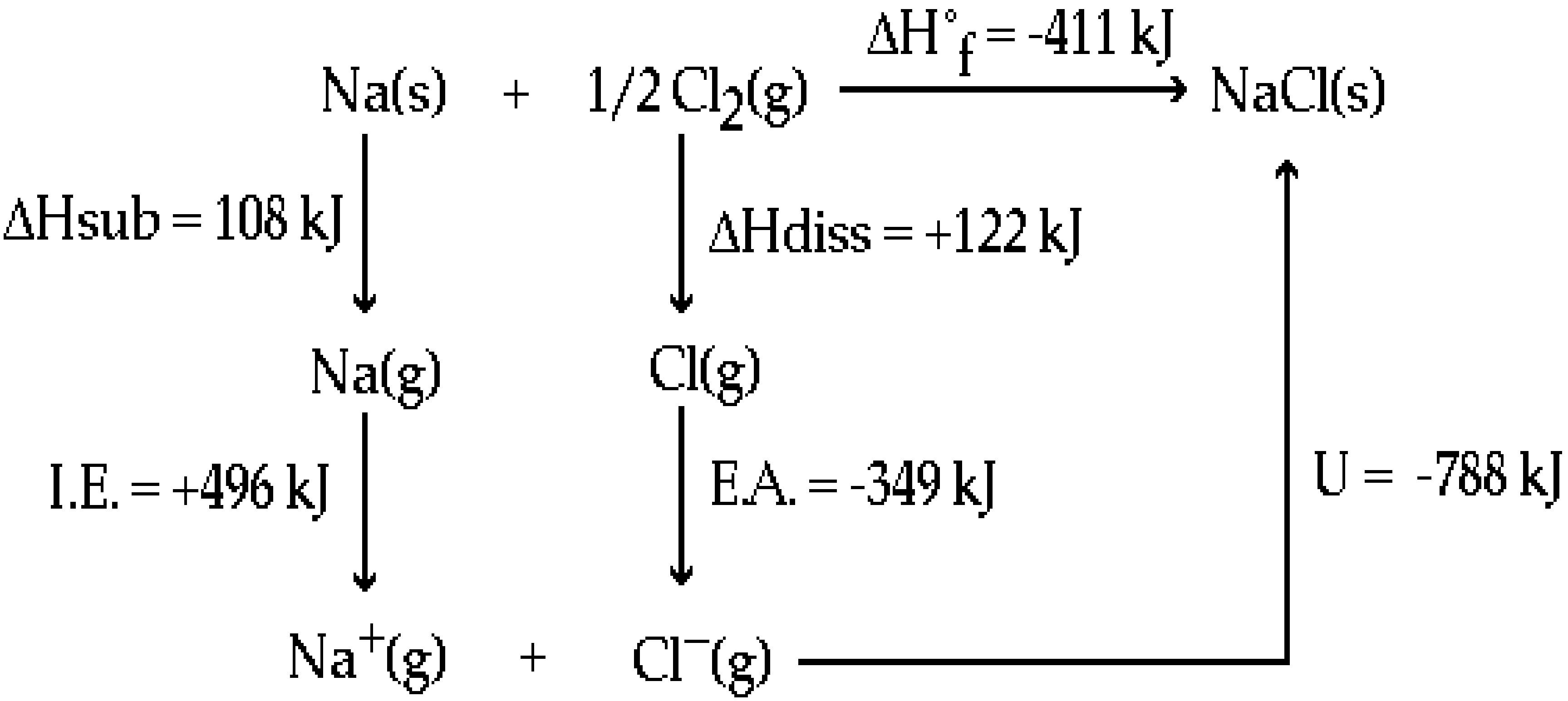
It is the electrostatic attraction between the oppositely charged ions that constitutes the ionic or electrovalent bond.

Factors Affecting Ionic Bond Formation

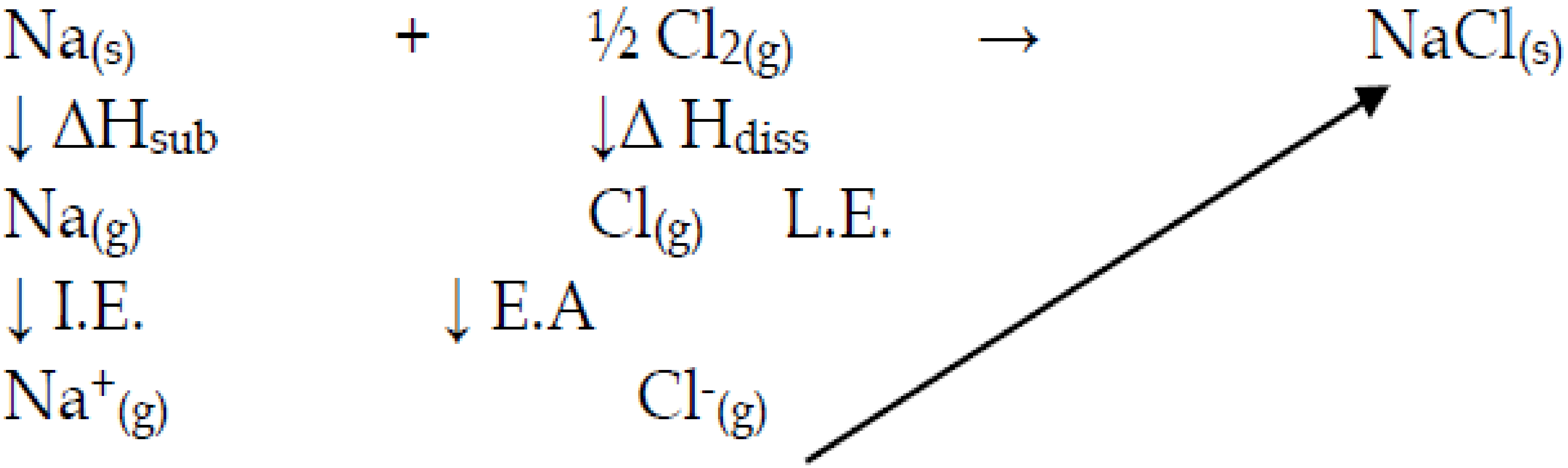
- The Born-Haber cycle can be used to illustrate the energy changes that take place when Na reacts with Cl to form NaCl.

- It is possible to form NaCl by the direct reaction between Na and Cl.

- The enthalpy change when one mole of a compound is formed from its elements in their standard states is the standard enthalpy of formation.



- It is also possible to form NaCl through several steps as shown in the cycle below.



$Na_{(s)}$ + $\frac{1}{2}\, Cl_{2(g)}$ → $NaCl_{(s)}$

↓ ΔH_{sub} ↓ ΔH_{diss}

$Na_{(g)}$ $Cl_{(g)}$ L.E.

↓ I.E. ↓ E.A

$Na^{+}_{(g)}$ $Cl^{-}_{(g)}$

ΔH_{sub} = enthalpy of sublimation

I.E. = Ionization energy

ΔH_{diss} = bond dissociation energy

E.A. = electron affinity

L.E. = Lattice energy

- Using the Hess' law of constant heat summation, it can be said that.

$$\Delta H_f = (\Delta H_{sub} + L.E.)Na + (\tfrac{1}{2}\Delta H_{diss} + E.A.)Cl + L.E.$$

- Taking the signs of the various energy changes into consideration, the equation can be rewritten as .

$$\Delta H_f = (\Delta H_{sub} + L.E.)Na + (\tfrac{1}{2}\Delta H_{diss} - E.A.)Cl - L.E.$$

The more negative the enthalpy of formation, the more stable the compound, or the more ionic the compound.

The three most important factors that will favour ionic bond formation (by making ΔH_f more negative) are;

- low ionization energy (low positive value)
- high electron affinity (high negative value)
- lattice energy (high negative value)

By way of summary, in the formation of a binary ionic compound, the atom which forms the cation should have low ionization energy:

The atom which forms the anion should have high electron affinity:

Then the interaction between the gaseous oppositely charged ions to form the crystal lattice should be so great that a large amount of lattice energy is released.

- An atom of low ionization energy has low electronegativity while an atom of high electronegativity.

- Thus ionic bond formation element favoured when the electronegativity difference high.

- This is the case for metals to the left side of the periodic table and non-metals which are to the right side.

- If the electronegativity difference between two elements that form a compound is above 1.7, the compound would be ionic.

- If it is slightly less than this, the compound would be polar covalent. But if the difference is very low, the bond would be non-polar covalent.

- Let us use the following atoms and their electronegativities to predict which pair of atom would form ionic bonds or otherwise.

Na (1.0) Al (1.5) O (3.5) Cl(2.8) F(4.0) P(2.1) H(2.2)

Atoms	Electronegativity difference
Na, O	3.5 - 1.0 = 2.5
Al, F	4.0 - 1.5 = 2.5
Al, Cl	2.8 - 1.5 = 1.3
O, F	4.0 - 3.5 = 0.5
P,H	2.2 - 2.1 = 0.1

The pairs of atoms with the largest electronegativity difference are (Na, O) and (AL, F) These would form the most ionic compounds.

The pair, (Al, Cl) forms covalent bond with the greatest ionic character.

The (P, H) has the least electronegativity difference, so the bond covalent with the least ionic character.

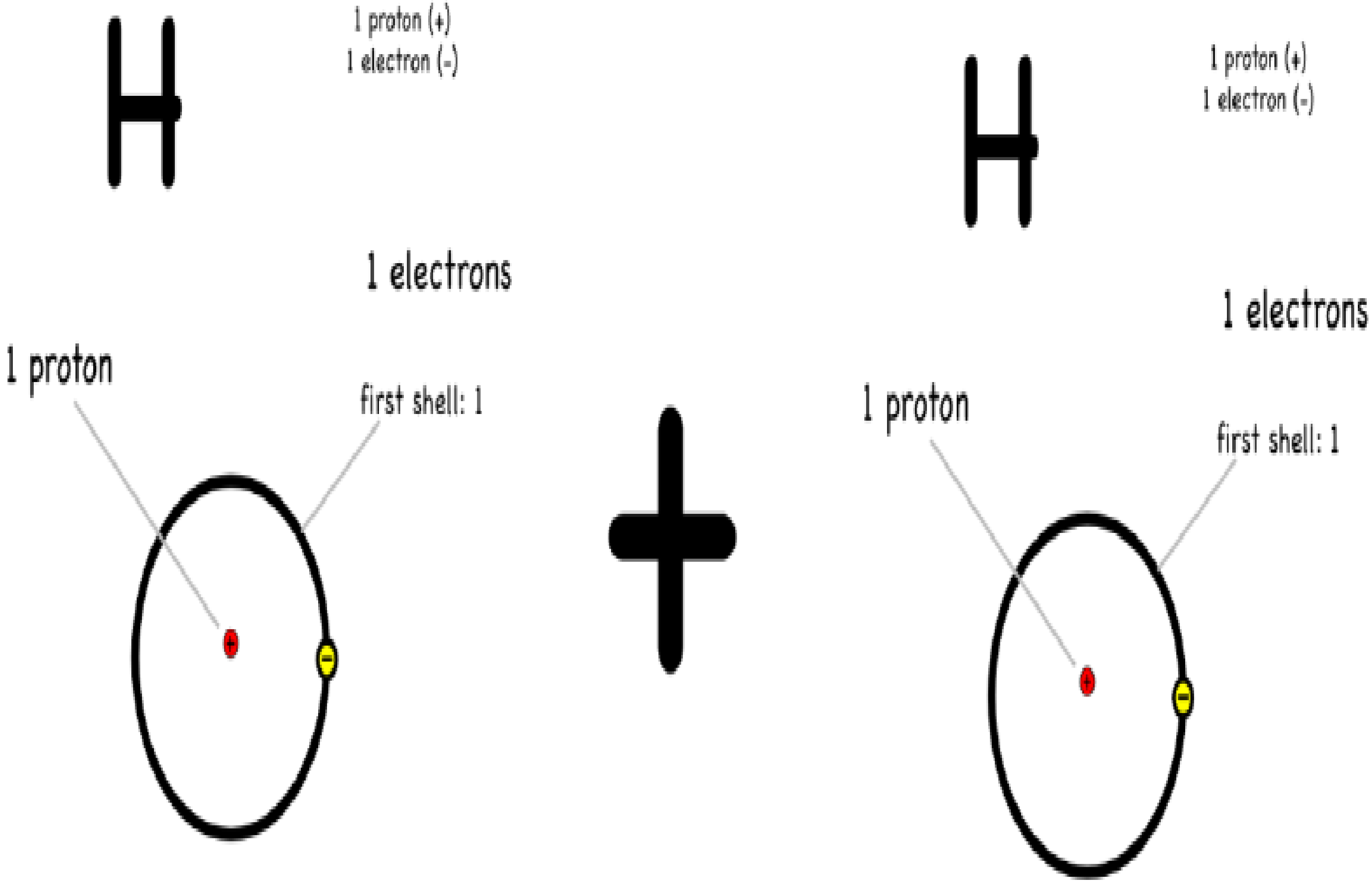
Atoms with the same or almost the same electronegativity cannot be bonded together through electron transfer.

They rather share electrons to form covalent bonds.

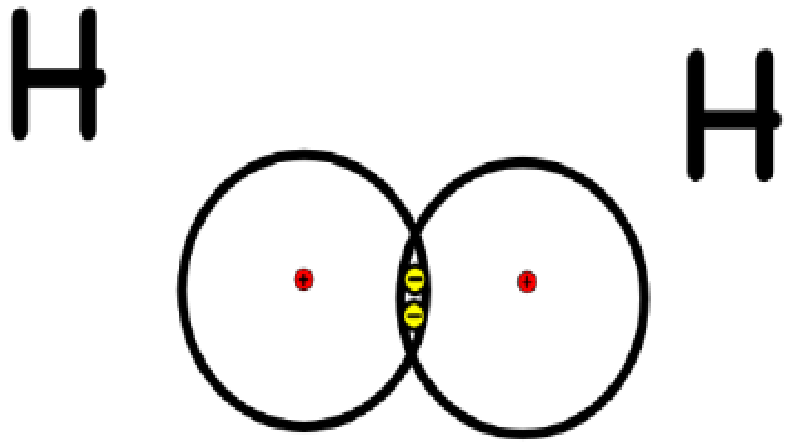
Each of the atoms involved in this type of bond formation donates the same number of valences electrons which are then shared between the bonded atoms.

Take the simple case of two hydrogen atoms combining, as each of the atom donates its valence electron, and the pair of donated electrons are then shared between the two nuclei, giving the hydrogen molecule, H_2.

- Covalent bond formation can be illustrated as follows:
- H + H → H:H
- H + H → H:HE


H
1 proton (+)
1 electron (-)
1 electrons
1 proton
first shell: 1
+
H
1 proton (+)
1 electron (-)
1 electrons
1 proton
first shell: 1

The two hydrogens share their electons.



Chemical Formula: H_2

- When two or more atoms are bonded covalently, the resulting particle is called a molecule.

- The attraction of the nuclei for the shared electrons forms the covalent bond.

- Where each of the bonded atoms donates an electron each as seen in the case of the hydrogen molecule, a single covalent bond is formed.

But where each of the bonded atoms donates two electrons which are then shared (two electron pairs are shared), double bond is an example of a molecule with a double bond is oxygen, O=O.

Where each atom donates three electrons which are then shared, a triple bond is formed as in N=N molecule.

Sigma and pi covalent bonds

- We will now look at two types of covalent bond; Sigma and pi bonds.

- In CH4 there are four C H bonds which are equivalent in energy and direction. In ethane $CH_2{=}CH_2$ apart from the C-H bond, the two carbon atoms are linked by double bond.

- The two bonds linking the carbon atoms are not equivalent. One is sigma covalent bond, (like the C-H bonds in CH_4) and other is pi-bond.

A sigma bond is formed when there is an end to end overlap of atomic orbitals to form a bond in which the electron density is concentrated in the region between two bonded nuclei.

On the other hand, a pi-bond is formed through the sideways or lateral overlap of orbitals. In a pi-bond the electron density alongside the molecular axis.

Because the head –on or end to end overlap of orbitals is more effective than the lateral overlap, sigma bond is stronger than pi-bond.

- It should be understandable that the C=C bond energy is not two times the C-C bond energy.

- However, the presence of the pi-bond makes the alkenes more reactive than the alkane.

- The pi-bond can be broken easily to make the electron of the available for reaction

- Hybridisation is the mutual mixing of atomic orbitals of different shapes and energy to form orbitals of equivalent energy, shape and direction.

- It has been mentioned in the last section that the four C-H bonds in CH_4 are equivalent.

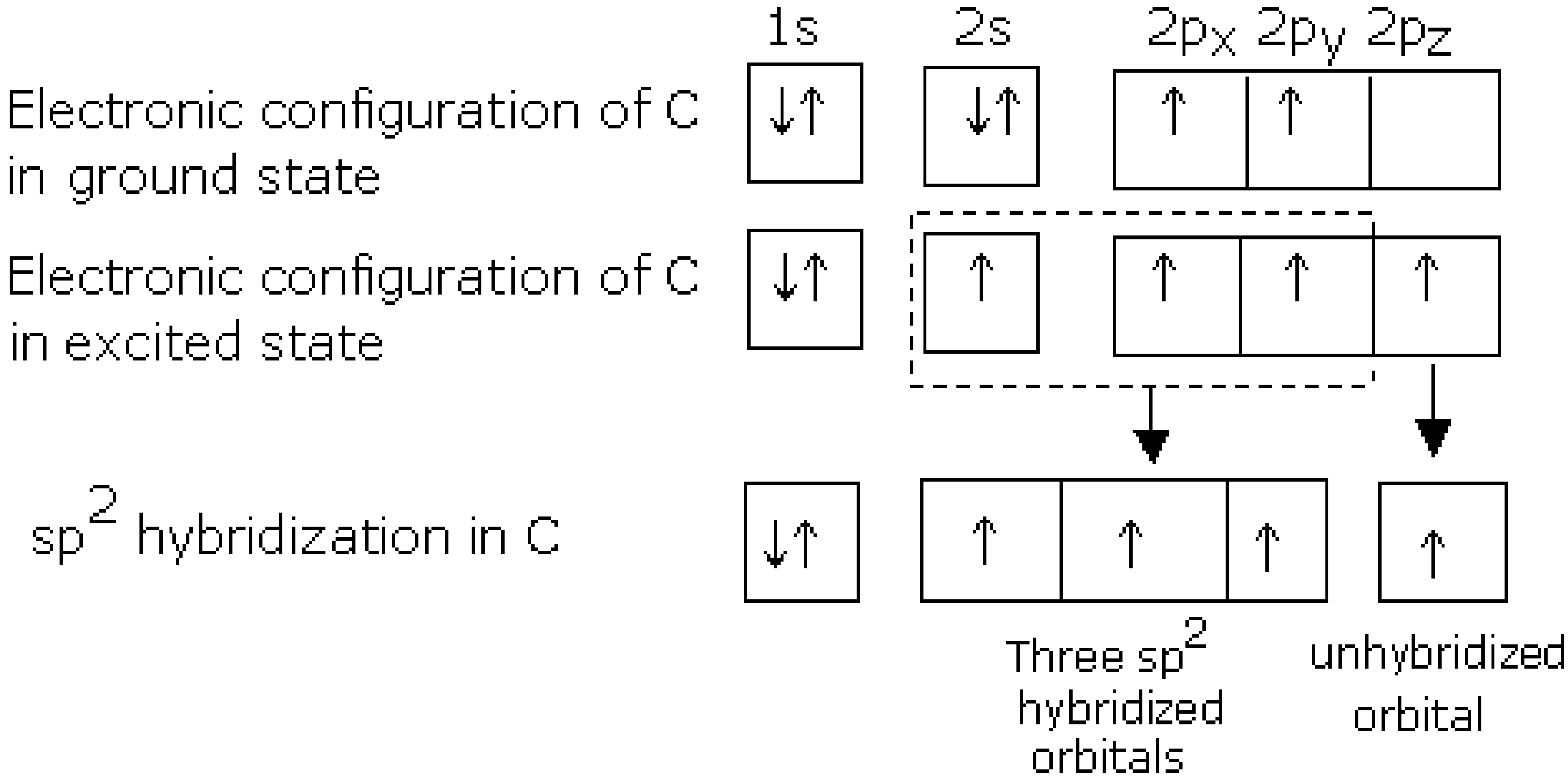
- Looking at the ground state electronic configuration of C, $1s^2\ 2s^2\ 2p^2$ we should have expected carbon to be divalent as there are only two unpaired electrons.

- But carbon is tetravalent.

- The tetravalency suggests that one of the valence electrons in the 2s orbital promoted to the 2p-orbital.

- This is what is supposed to happen in the excited state of carbon prior to its formation of compounds.

The electronic configuration of the excited state of carbon will be



1s 2s $2p_x$ $2p_y$ $2p_z$

Electronic configuration of C in ground state ↓↑ ↓↑ ↑ ↑

Electronic configuration of C in excited state ↓↑ ↑ ↑ ↑ ↑

sp^2 hybridization in C ↓↑ ↑ ↑ ↑ ↑

Three sp^2 hybridized orbitals

unhybridized orbital

- Now there are four unpaired valence electrons, and so the tetravalency of carbon is accounted for.

- However, there is yet another outstanding: the fact that the bonds formed by C are equivalent in energy.

- The four valence electrons are in two different orbitals, 2s and 2p.

- The s-orbital is spherical while the p-orbital is dumb-bell in shape, so they are not expected to form bonds which are equivalent.

- It is to account for the equivalence of bonds formed from different atomic orbitals that the concept of hybridization has been proposed by some chemists.

- There can be the mixing of one 2s-orbital and the three 2p orbitals to form sp3 hybrid orbitals which is supposed to be at the center of a tetrahedron.

- The hybrid orbitals of the same size, shape; and the angle between any two of these orbitals is tetrahedral (109^0).

- It is the hybrid orbitals which are presumed to overlap with s-orbitals of four hydrogen atoms to form methane (CH_4) with the four equivalent C-H bonds.

- In ethane, the carbon atoms are sp3-hybridized.

- This means that in the excited state of carbon the one 2s-orbital mutually mixes with two of the 3p-orbitals forming the sp3-hybrid orbitals with an angle of 1200 between any two hybrid orbitals.

- The shape of the sp2-hybrid orbitals is planar or flat.

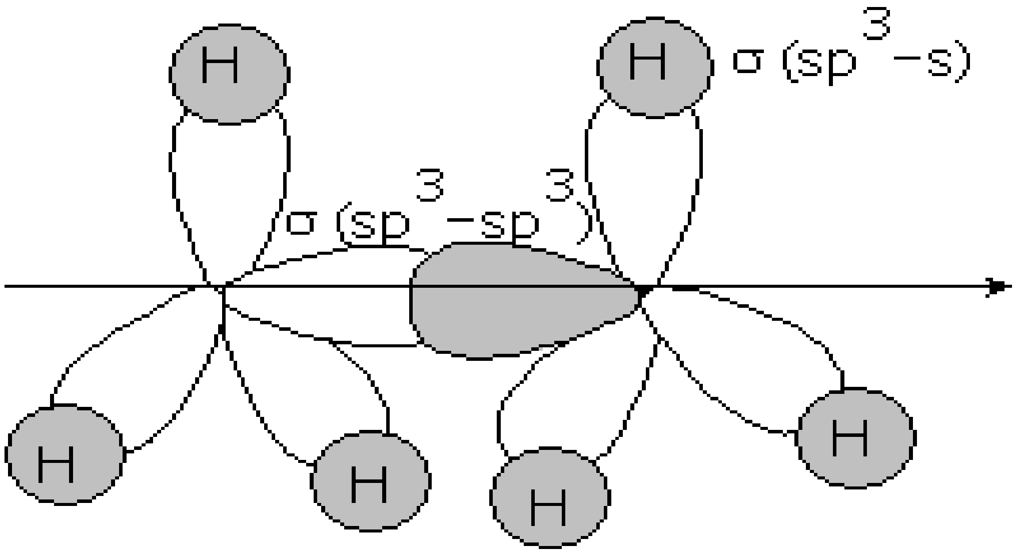
- Each carbon uses one of the hybrid orbitals in overlapping with the other in an end to end manner to form a C-C sigma covalent bond.

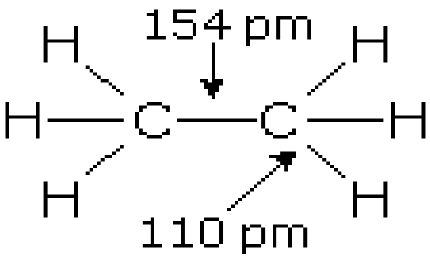
- The two remaining hybrid orbitals on each carbon overlap with two s-orbitals of hydrogen.

- Left on each carbon is one unhybridized p-orbital.

The two unhybridized p-orbitals are parallel to each other, and also perpendicular to the plane of the compound.

There is a lateral or sideways overlap of these unhybridized p-orbitals to form a pi-bond.





- Having come this far, make an attempt to work out the structure of ethyne, H-C≡C-H.

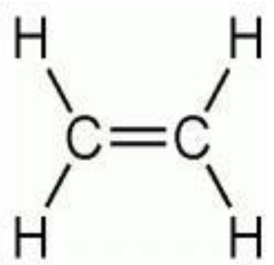
- Note the following:

- Type of hybridisation of carbon atoms –sp

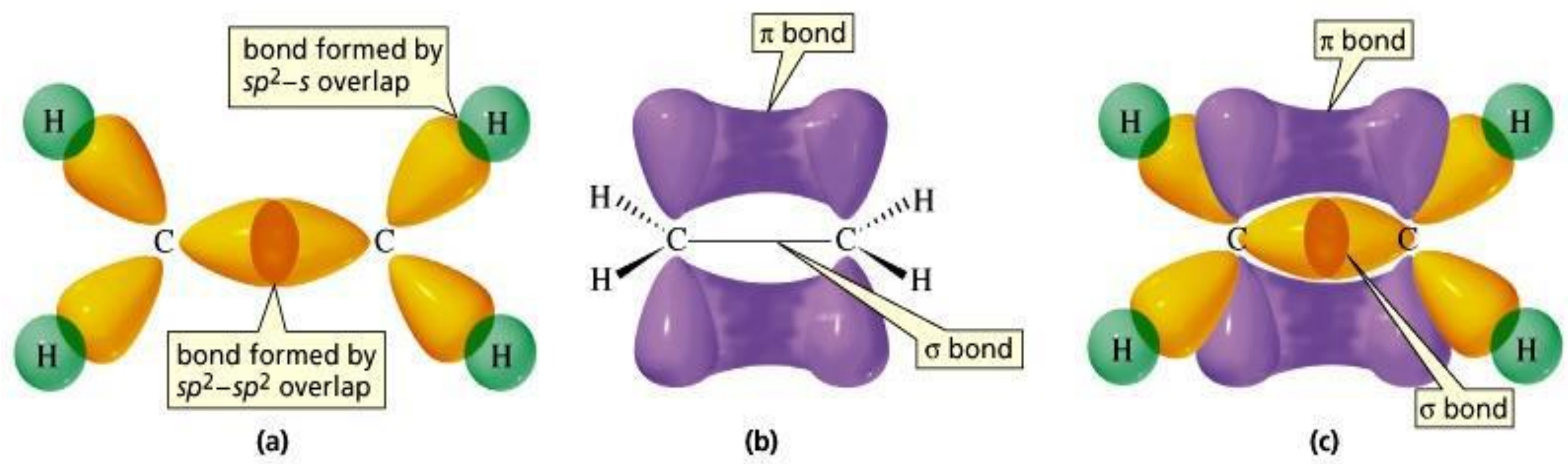
- Number of hybrid orbitals – two

- Arrangement of hybrid orbitals about carbon – linear.
- Angle between hybrid orbitals – 180^0 (dihedral).
- Number of unhybridised p-orbitals on each carbon – 2.
- Types of bonds linking two carbon atoms – one sigma and two pi-bonds.

- Consider the hydridization in ethylene;



Ethylene (Ethene)



Coordinate or dative covalent bonds

- A coordinate bond is similar to a covalent bond in the sense that electrons are shared between bonded atoms.

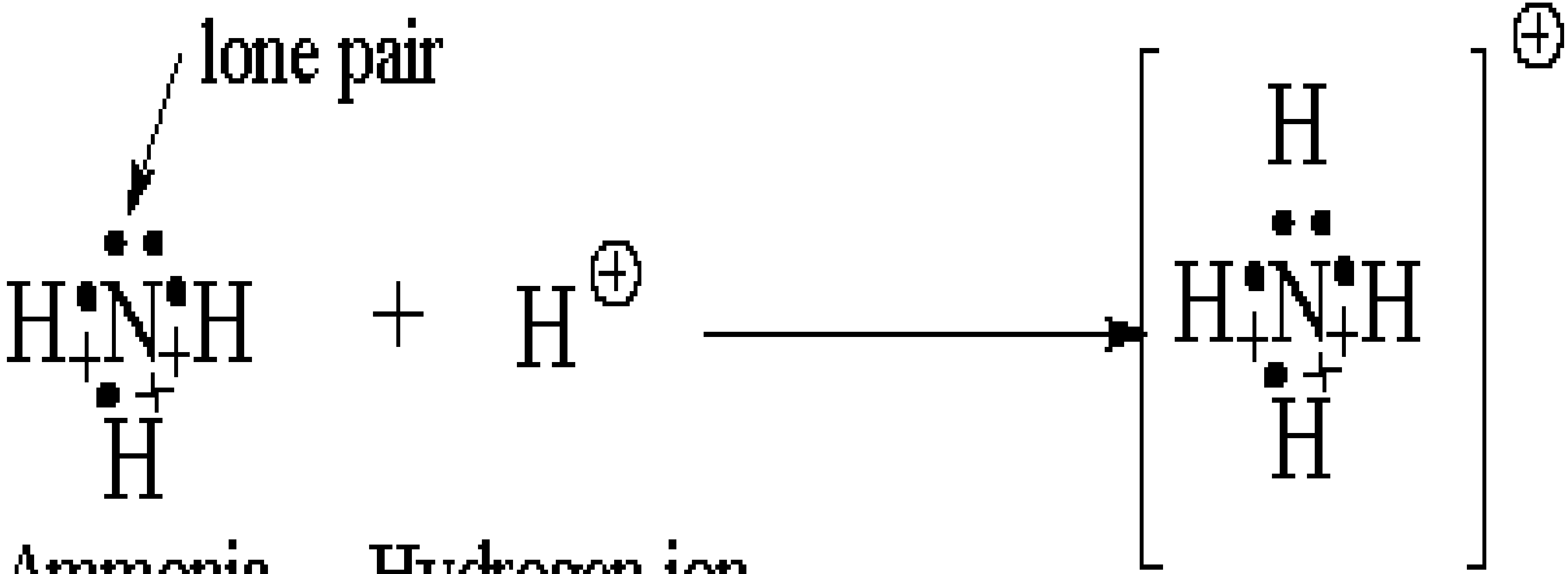
- However, in a coordinate bond formation, the shared electron pair/pairs are donated by one of the bonded atoms.

- So in this type of bond formation, there should be the donor atom which will donate a pair of electrons, and the acceptor atom which should have empty orbital.

- A simple example of a species which has a lone pair of electrons is Ammonium ion NH_4^+.

- Ammonia (NH_3) has a nitrogen atom which has a lone pair of electrons. Hydrogen ion (H^+) is an acceptor; it has an empty s-orbital because the only electron in the hydrogen atom has been lost.

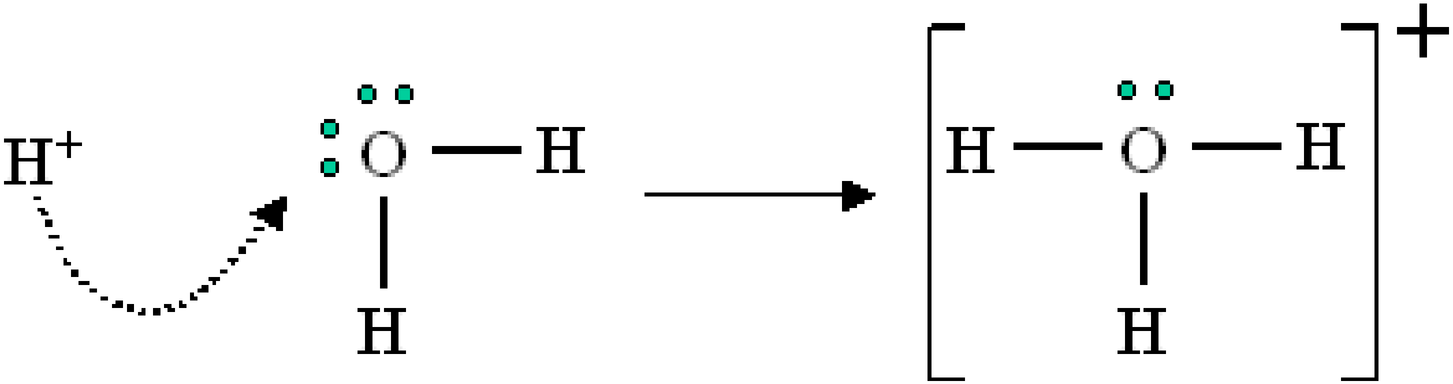
- The lone pair of electrons on the nitrogen is donated to the H^+, and this electron pair is now shared between the N and H, giving rise to the new species, NH_4^+.



Ammonia + Hydrogen ion → NH_4^+

The ammonium ion

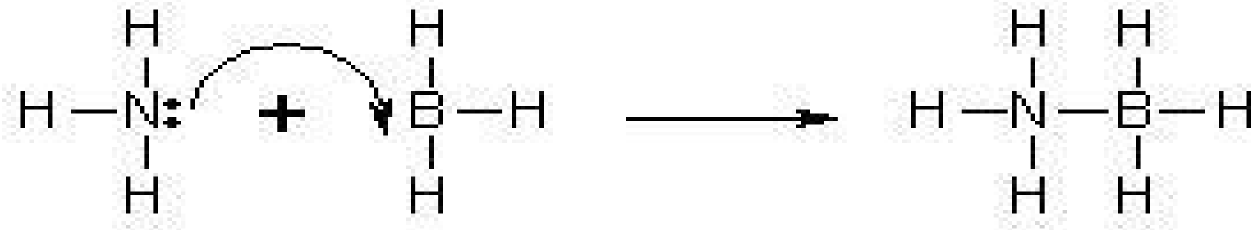
The oxonium ion, H_3O^+ also contains this type of bond as water H_2O, by virtue of the presence of lone pairs of electrons on its oxygen, can donate a lone pair to a hydrogen ion to form an oxonium ion, H_3O^+.



- NH3 can also react with BF3 to form a compound with coordinate bond.

- The donor atom, N in NH3 can donate a lone pair of electrons, and the acceptor atom.

- B in BF3 has an empty orbital, so can accept the electron.
- $BF_3 + NH_3 \rightarrow H_3N\text{-}BF_3$



- The electron pair donors could be molecules or they could be anions.

- For the acceptors, they could be electron-deficient cations or compounds.

- In transition metal chemistry, coordination complexes are formed in which ligands with lone pair of electrons donate electron pairs to metallic coordination centers to form coordinate bonds.

- Examples of such complexes which could be cationic, neutral or anionic are $[Cu(NH_3)_4]_2^+$ and $[Fe(CN)_6]_4^-$.

- Some important organic molecules in the body contain coordinate bonds.

- Haemoglobin, the oxygen carrier in the blood of mammals has a haem group in which Fe_2^+ forms dative covalent with the nitrogen of some pyrrole rings.

- According to the octet rule put forward by Lewis, in compound formation, atoms transfer or share electrons in order to attain a noble gas configuration with an octet.

- Indeed, the essence of chemical reactions is for the elements involved in the reaction to attain a noble gas electronic configuration which is a stable arrangement.

- In the reaction between sodium and chlorine to form NaCl, the Na^+ as well as the Cl^- in the compound attain the noble gas configuration of Ne and Ar respectively.

- The octet rule holds for some period 2 elements like carbon, nitrogen, oxygen and fluorine.

- But it fails for elements in which d-orbitals are available.

- For example, in PCl_5, the central atom, P uses its d-orbitals to accommodate an additional electron pair beyond the octet.

- If a molecule has an atom in which the valence shell contains more than eight electrons as in the case of PCl_5, the central atom of the molecule is said to have an expanded octet.

- In the last section, we came across BF; which was said to be electron-deficient.

- The outermost shell of boron in this compound has six electrons.

- Here too, the octet rule is not obeyed.

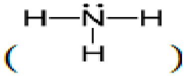
Bonded versus lone pair of electrons

- In the formation of molecules, the uncombined atoms should have unpaired electrons in each orbital of the outermost shell.

- In the formation of covalent bonds, these unpaired electrons then pair up.

- Therefore, bonded electrons are the pair of electrons in the valence shell that belong to the atoms that are bonded together in a molecule.

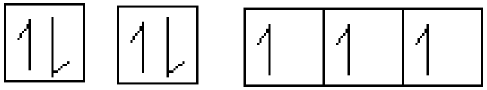
- Lone pair of electrons or unshared electron pairs are also found in the valence shell of an atom in a molecule, but the lone pair belongs to only one of the bonded atoms.

- In NH3, the central atom, N is bonded to three hydrogen atoms  Each N-H bond is made up of a pair of electrons from the two bonded atoms.

- In addition, the outermost shell of N contains a pair of electrons which is exclusively for the N.

- This is the lone pair of electron or the unbonded electron pair.

- Let us recall the electronic configuration of N so that we can appreciate the presence or the bonded and non-bonded electron pairs in NH3.

- N = $ls^{2}2s^{2}2p^{3}$ or 

- The valence electrons in N which can be used for bond formation are the 2s and 2p electrons.

- When hydrogen combines with nitrogen to form ammonia, it can simply be as the three s-orbitals of hydrogen overlapping with the three unpaired electrons to form the three N-H bonds.

- These are the bonded electron pairs.

- The other paired valence electron on the N now become the lone pair in NH_3.

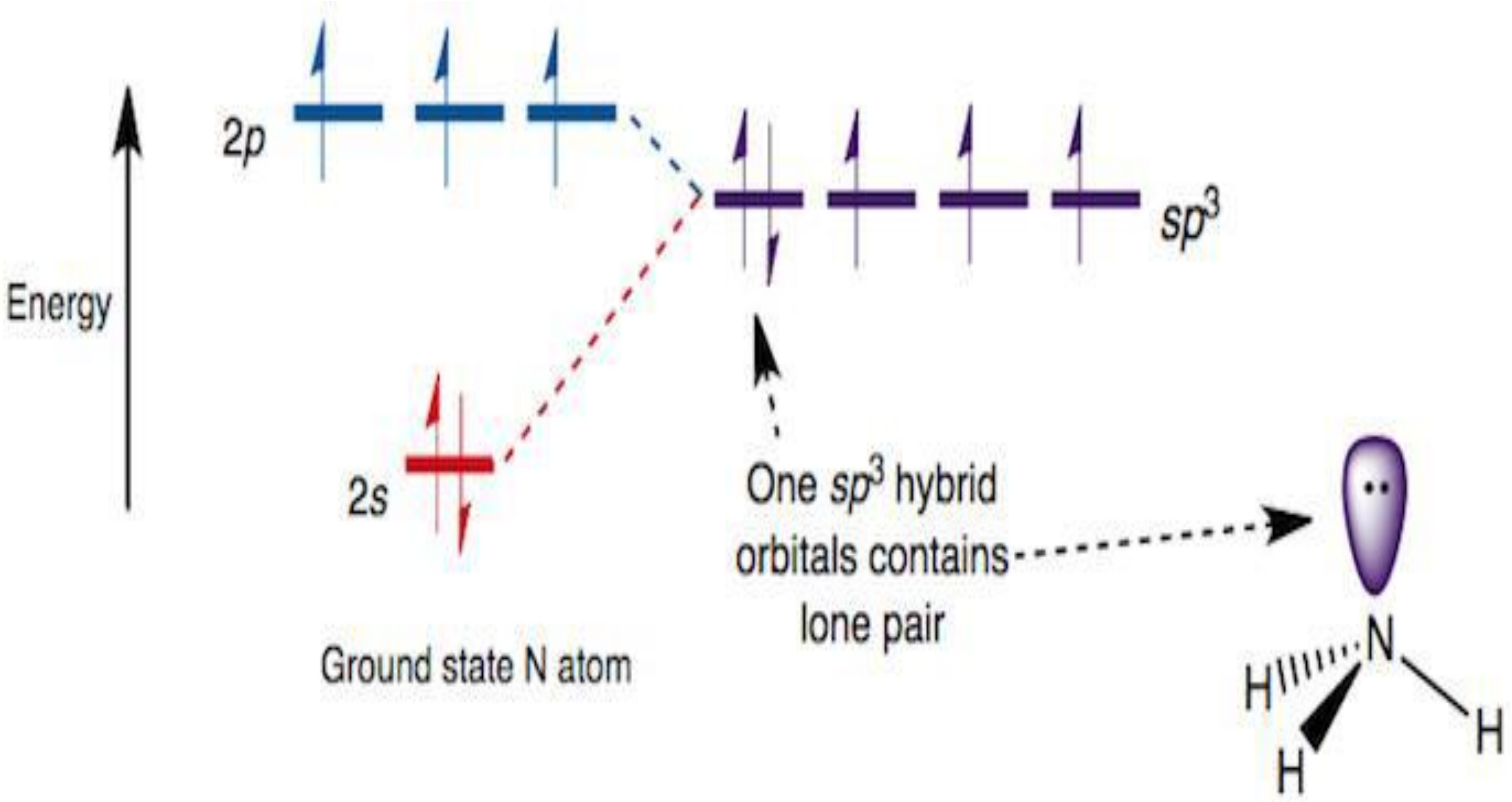
- As an isolated atom, N does not have a lone pair.

- Because the lone pair belongs to only one atom in a molecule, it provides a region of greater electron density than a bonded pair in which the electron density is spread between the two bonded atoms.

- Furthermore, a lone pair in a molecule is readily donated.

- For example, in the coordinate bond formation, it is the readily available lone pairs of the donor atoms which are supplied to acceptors.

- In the same way, Lewis bases like NH3 and OH- behave the way they do because they can make a lone pair readily available.


2p
sp^3
Energy
2s
One sp^3 hybrid orbitals contains lone pair
Ground state N atom
N
H
H
H

Shapes of molecules

- The shape of a molecule is an important property of the molecule because it affects the behaviour of the molecule.

- The shapes of many organic molecules determine their reactivity odour, taste and their action as drugs, in some cases.

- The interaction between enzymes and substrates depends on their shapes.

- There is what is called the ‘lock and key‘ hypothesis of enzyme-substrate interaction.

- The enzyme is supposed to be a lock with a specific shape, so it is only a complementary substrate (key) with the appropriate shape that can fit into the lock.

- This explains why enzymes are more specific in their catalysis.

- Hormones are able to act because these chemical messengers, with characteristic shape, depending on the hormone, can interact with some receptors, also of specific shape in order to elicit some physiological response Molecules, as they are formed, contain some electron pair in their valence shells.

- As pointed out in the last section, these electron pairs could be bonded or unbonded.

- Because these electron pairs are centers of negative charges, they would tend to repel.

- The arrangement of these electron pairs should be such as to minimize repulsion in order to ensure the stability of the molecules.

- This idea was suggested by Sidgwick and Powell and it is called the valance shell electron pair repulsion (VSEPR) theory.

- Another noteworthy point is that the repulsions in the valence electron pairs are not the same.

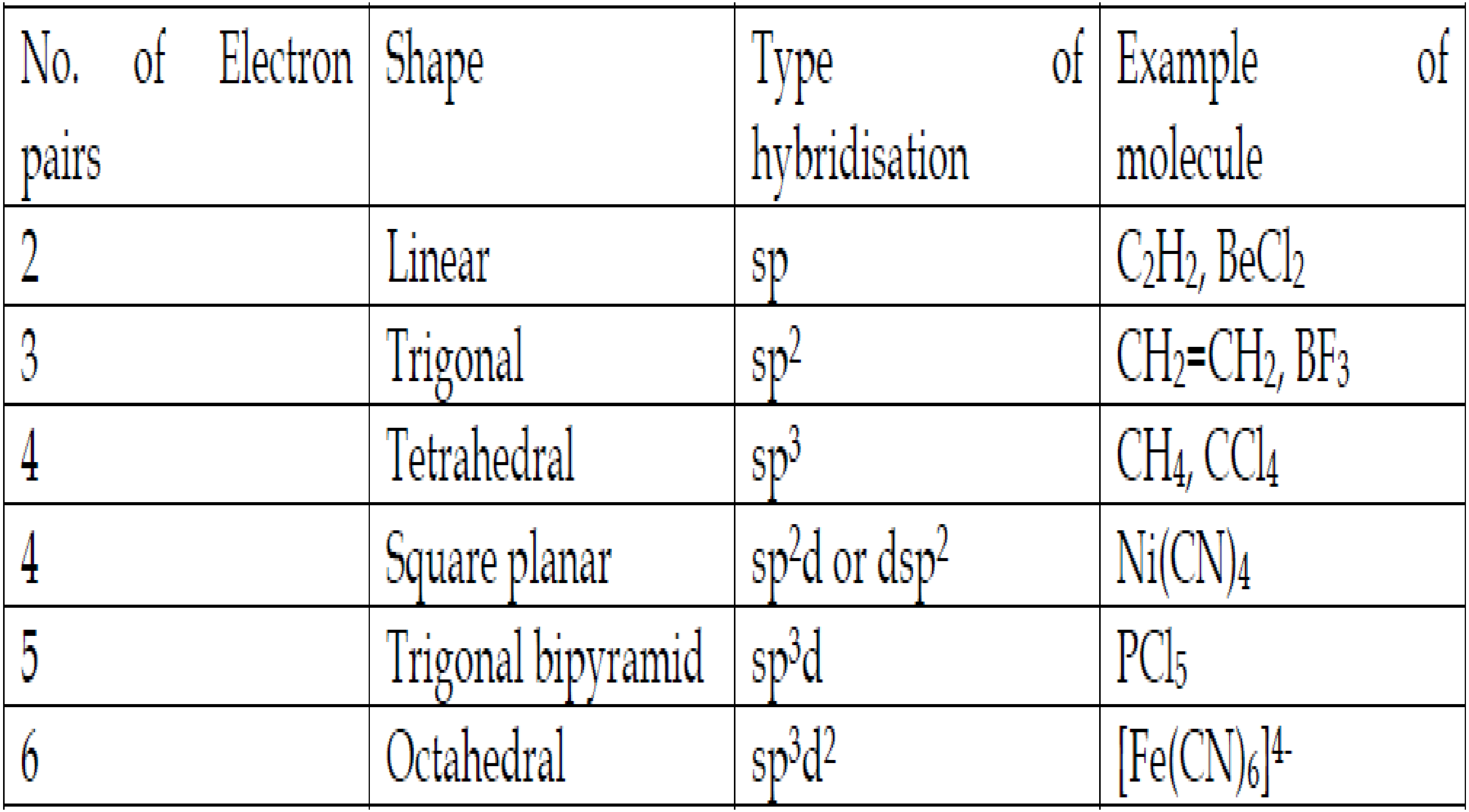
- The repulsion between lone pair-lone pair electrons is greater than the repulsion between lone pair and bonded pair, while the repulsion between bonded pair — bonded pair would be the least.

- Given any molecule, the central atom should be identified.

- Then the electron pairs surrounding this atom should be worked out in order to find out the shape of the molecule.

- From the shape of the molecule, the type of hybridization in conformity with the shape can be deduced.

The table below is an illustration of what has so far been discussed.



No. of Electron pairs	Shape	Type of hybridisation	Example of molecule
2	Linear	sp	C_2H_2, $BeCl_2$
3	Trigonal	sp^2	CH_2=CH_2, BF_3
4	Tetrahedral	sp^3	CH_4, CCl_4
4	Square planar	sp^2d or dsp^2	$Ni(CN)_4$
5	Trigonal bipyramid	sp^3d	PCl_5
6	Octahedral	sp^3d^2	$[Fe(CN)_6]^{4-}$

Bond angles

- Unlike ionic bonds which are extended in all directions, covalent bonds are direction.
- Any two bonds are med by an angle, depending on the shape of the molecule.
- We have already seen that the bond angle in methane in which the central carbon atom is sp3-hybridised is 109°.
- Then, the bond angle in ethane is 120°.

- In NH3, the central atom is N which has four valence electron pairs surrounding it, like in C of methane.

- The shape of NH3 should be tetrahedral, and the type of hybridization is sp3, we should also expect the tetrahedral bond angle.

- But we should not forget that of the four valence electron pairs around nitrogen, one of them is a lone pair.

- The repulsion between the lone pair and bonded pairs will tend to decrease the tetrahedral angle.

- This explains why the bond angle of NH_3 is 107°.

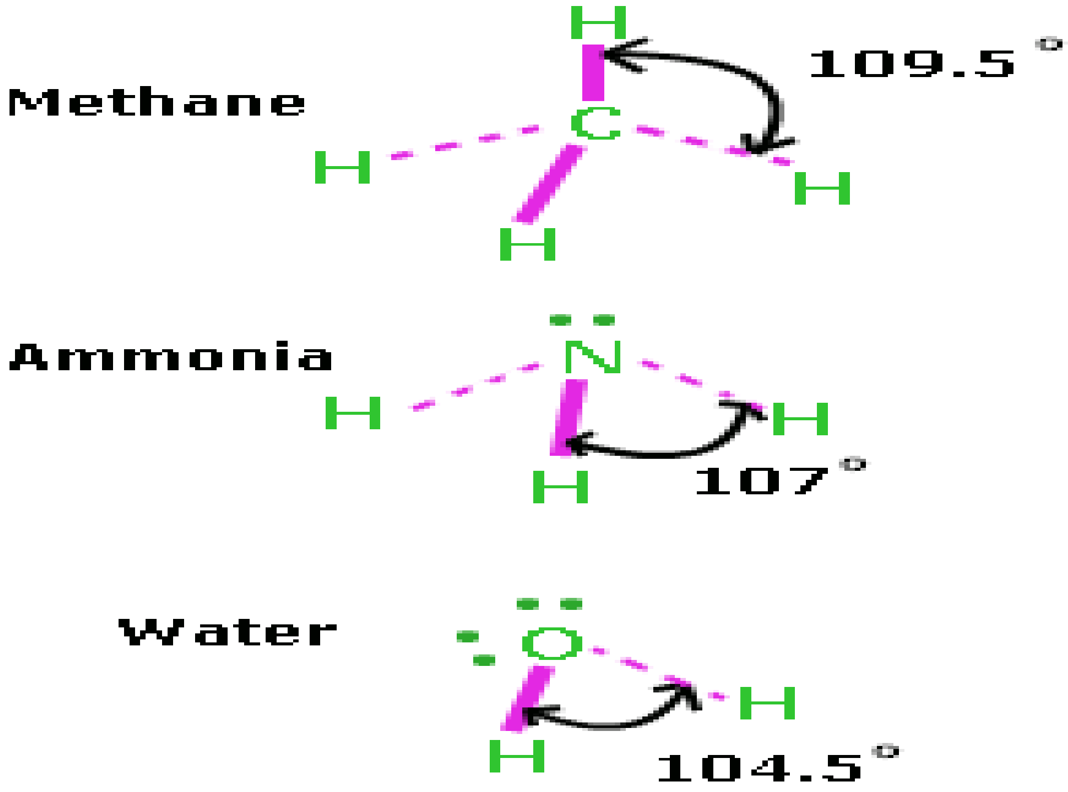
- Water, H_2O has two bonded pairs and two lone pairs on oxygen, its central atom.

- The repulsion between the two lone pairs will even be greater, and this reduces the bond angle to about 104.5^0.

- As a result of the repulsion due to the lone pair of electrons in

- NH_3 and H_2O, the two molecules are said to have distorted tetrahedron shapes.

- The actual shape of NH_3 is trigonal pyramid while that of water is angular or v-shaped.



APPREY CHARLES

Polar and non-polar molecules

- Covalent molecules could be formed by atoms which are similar or different.

- Hydrogen or molecules (H_2, Cl_2) are made up of atoms. In these two molecules, none of the bonded atoms has the tendency to attract the shared electron pair.

- Such molecules are said to be non-polar: the two bonded atoms have the same electronegativity.

- If we consider HCl molecule, we see a different situation.

- The two bonded atoms have different electronegativities. The Cl is more electronegative than H.

- By virtue of its higher electronegativity, the Cl attracts the shared electron pair to assume a partial negative charge while the less electronegative H, becomes partially positive charged: $H^{\delta+} – Cl^{\delta-}$

- The HCl is said to be a polar molecule ‘the greater the electronegativity difference, the more polar the molecule.

- For the hydrogen halides, HF, HCl, HBr and HI, we expect HF to be the most polar while HI would be the least polar.

- It is not only the electronegativity difference that determines the polarity of molecules, the shape of molecules is another factor CO_2 is non-polar.

- This is a linear molecule, O=C=O, made up of two CO bonds which are polar due to the electronegativity difference between C and O.

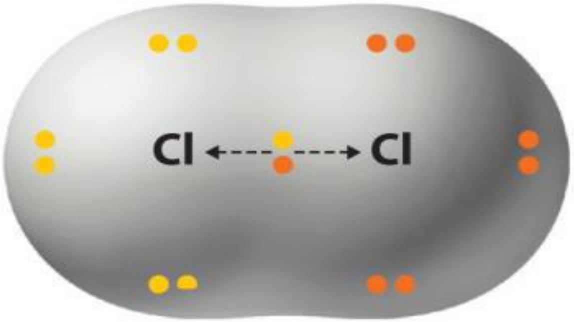
- However because of the linear nature of CO_2, the polarities cancel making the overall molecule non polar.

- CCl_4 is non-polar, even though it contains individual C-Cl bonds which are polar due to the electronegativity difference between C and Cl.

- But in this case due to the tetrahedral structure, the CCl_4 is symmetrical, so the polarities cancel.

- A molecule like HCl has a positive charge center as well as a negative charge center.

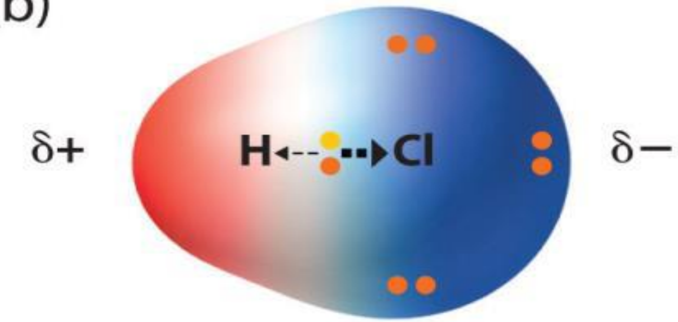
(a)



Nonpolar covalent bond

Bonding electrons shared
equally between two atoms.
No charges on atoms.

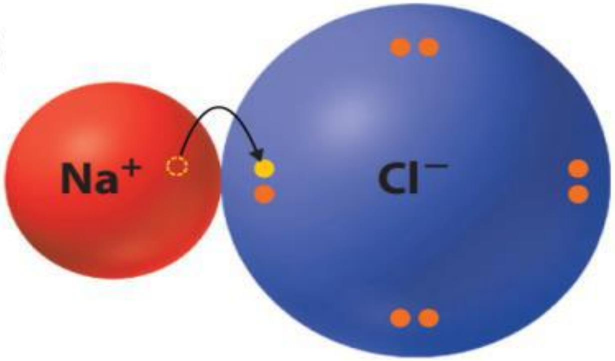
(b)



Polar covalent bond

Bonding electrons shared
unequally between two atoms.
Partial charges on atoms.

(c)



Ionic bond

Complete transfer of one or more
valence electrons.
Full charges on resulting ions.

- Such a molecule is called a dipole, or it is said to be dipolar.

- A dipole moment gives a quantitative measure of the polarity of a molecule.

- The dipole moment is equal to the product of the charge on the molecule (could he the positive or negative charge) and the distance of separation between the positive and negative charge centers.

- The unit used is the debye (D).

- The higher the dipole moment, the more polar the molecule.

- CO_2 and CH_4 have zero dipole moment, but H_2O, NH_3 and H_2S have net dipole moments.

- Where there is a lone pair of electrons in a molecule, a net dipole moment points towards the lone pair.

- Then for linear, trigonal, tetrahedral and octahedral molecules, due to the symmetrical arrangement of bonds, the molecules have zero dipole moment provided the central atom is bonded to only one type of atom/group.

- Do you expect $CHCl_3$ to be polar or non-polar? Explain

Intermolecular forces

- Covalently bonded compounds occur in all the physical states; gas, liquid and solid Methane (CH_4), hydrogen sulphide (H_2S) are gases, water (H_2O), chloroform ($CHCl_3$) are liquids while benzoic acid (C_6H_5COOH) is solid.

- The differences in the physical states depend on the intermolecular forces.

- If these forces are strong enough they keep the molecule in fixed position as found in the solid state.

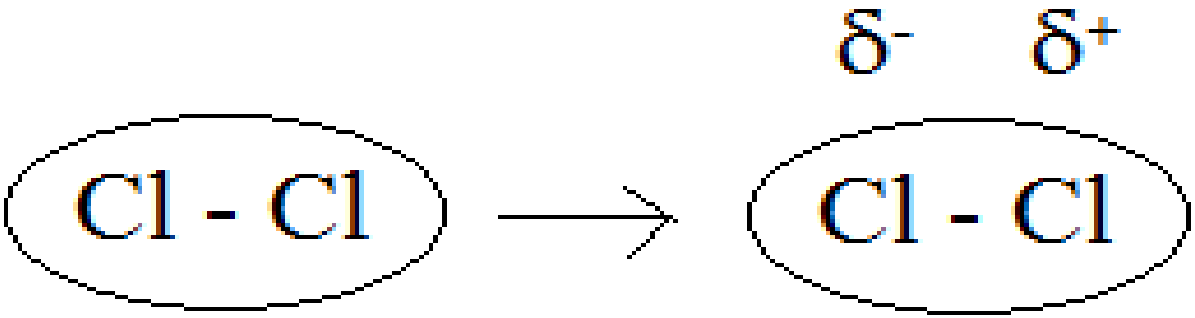
- These intermolecular forces could be permanent dipole-dipole attraction or temporary induced dipole-induced dipole attractions.

Van der Waals forces

- One type of intermolecular forces found in both polar and non-polar molecules is the Van der Waals forces or the London forces.

- They are instantaneous, short range, temporary induced dipole- induced dipole attractions.

- Even for non-polar molecules like Cl_4 or CCl_4, it does happen that at an instant the electrons cloud associated with the molecules become distorted, so that a temporary dipole is created.



- This dipole then induces the formation of another dipole in a neighbouring molecule.

- For example, the negative end of the molecule would repel the electron in a nearby molecule to induce the formation of a partial positive charge at one of that nearby molecule while the other end of the molecule would be partially negative charged

- The attraction between continuously charging instantaneous induced dipoles constitutes the van der Waals forces.

- The magnitude of these forces depends on the molecular size. The larger the molecular size, the more electrons there are in the molecule, and the greater the tendency for the distortion of the electron cloud to generate temporary dipoles.

- The fact that there is a gradation in the physical states of the halogens as the group is descended is a reflection of increasing van der Waals forces. F_2 and Cl_2 are gaseous, Br_2 is liquid, while I_2 is solid.

- The trend of boiling points in the halogen halides can be accounted for by the differences in the magnitude of the intermolecular forces.

- HF is a liquid of relatively high boiling point. But all the others, from HCl to HI are gases, with the boiling point increasing from HCl to HI as the molecular size increases in that order.

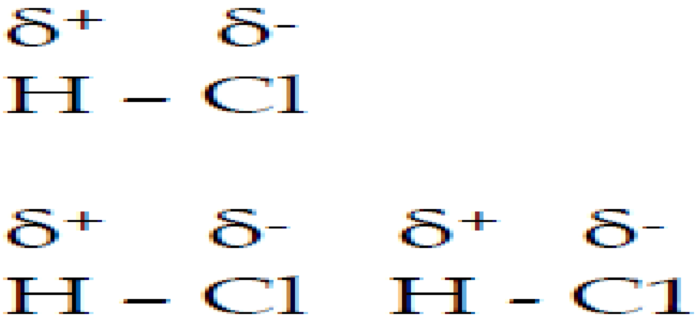
- Even though HF has the smallest size, due to the presence of hydrogen bonding, it has stronger intermolecular forces, making it liquid.

Permanent dipole-dipole attraction

- We have already seen that HCl is covalently bonded, but as a result of the electronegativity difference between the two atoms, the molecule is said to be polar, having partially positive and negative charges, $H^{\delta+}$-$Cl^{\delta-}$.

- An adjacent molecule of HCl would be aligned in such a way that the negative end of one molecule will be attracted to the positive end of another molecule.

- , e. g



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

$$\overset{\delta+}{H} - \overset{\delta-}{Cl} \quad \overset{\delta+}{H} - \overset{\delta-}{Cl}$$

- Such intermolecular forces are permanent dipole-dipole attraction.

- A special type of permanent dipole-dipole attraction is hydrogen bonding.

- Whenever hydrogen is covalently bonded to very small and electronegative atoms like fluorine, oxygen and nitrogen (FON), hydrogen bonding occurs.

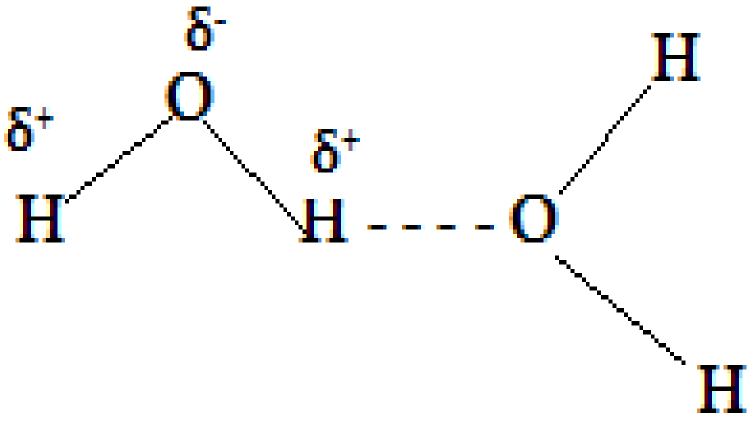
- Take the case of water, H_2O in which two hydrogen atoms are bonded to oxygen:

- Owing to the electronegativity differences between oxygen and hydrogen, the shared electron pair between the oxygen and hydrogen is drawn towards the oxygen,

- Making it partially negative charged while each of the hydrogen end becomes partially positive charged-Slight should not be lost of the fact that hydrogen atom has only one electron.

- So when the shared electron pair is withdrawn towards the oxygen, each hydrogen appears just like a bare proton.

- With no screening from any electron, the positive charge of the proton is so intense that if there is a nearby electronegative atom of an adjacent molecule with a lone pair of electrons, then these can be donated to the hydrogen to form another bond, which is the hydrogen bond.



- This type of bond is represented by broken lines, unlike a covalent bond which is indicated by a solid line.

- The hydrogen bond is weaker than a covalent bond.

- The presence of hydrogen bond allows a hydrogen atom to form a bridge between two electronegative atoms of two molecules, as shown in the illustration above.

Effects produced by intermolecular forces

- Intermolecular forces are not as strong as intramolecular covalent bonds.

- For this reason, the intermolecular forces do not directly affect the chemical properties of compounds; they rather affect physical properties.

- Such properties include physical states of compounds, solubility boiling point, melting point, specific heat, molar heat of fusion and molar heat of vaporization.

- Others are viscosity and surface tension HF behaves differently from the other halogen acids, HCl, HBr and HI. HF is liquid whereas the others are gases.

- Hydrogen bonding is possible in HF, slightly possible in HCl, non-existent in HBr and HI.

- The strong hydrogen bonding in HF results in the associated nature of the molecules, accounting for the liquid nature. In the same way, water (H_2O) behaves differently from the hydrides of the other group Vl elements.

- H_2O is liquid, but H_2S, H_2Se, H_2Te are gases.

- The H_2O molecules become associated through hydrogen bonding, but the other molecules have little or no hydrogen bonding, and so they occur as gases at room temperature.

Effect of hydrogen bonding on solubility

- There is this maxim ‘like dissolves like’.
- For example, ionic compounds are dissolved in polar solvents but do not dissolve in non-polar solvents.
- NaCl dissolves in water but not in ether.

- For a solute to dissolve in a solvent, the attraction between the solute particles and the solvent should be greater than the attraction between the solute particles.

- Thus NaCl dissolves water because the attraction between the Na^+/Cl^- particles and water is greater than the concentration between the Na^+ and Cl^- ions.

- On the other hand NaCl does not dissolve in ether as the Na^+ and Cl^- ions cannot be attracted to the non-polar ether.

- Candle wax does not dissolve in water.

- The strong hydrogen bonding in water prevents any interaction with the hydrocarbon wax.

- However, the hydrocarbon wax will dissolve in petrol.

- Here the van der Waals attraction between the candle and petrol will be greater than the van der Waals attraction between the solute wax particles.

Other effects produced by hydrogen-bonding

- The high boiling point of water ensures that under normal conditions, water remains a liquid.

- If water were to be like volatile ether, it would have vapourised even at ambient temperatures.

- As a result, there would not have been any collection of bodies of water like lakes, rivers or oceans.

- There would not have been any rainfall, either.

- Water has a high heat capacity.

- This suggests that a lot of calories is required to raise a given mass of water by 1°C.

- The high heat capacity of water prevents large bodies of water from warming or cooling too rapidly.

- This slow warming and cooling afford protection of living organisms from abrupt temperature changes as climate changes are moderated.

- Water also has a very high heat of vapourisation.

- It requires 540 calories of heat to vapourise one gram of water at 100°C.

- This enables water to absorb and store heat.

- Our bodies take advantage of the high heat of vapourisation of water.

- Each gram of water that vapourises from human skin carries away more than 500 calories of heat.

- Therefore, as we sweat, large amounts of heat are removed, producing a cooling effect when the weather is Warm.

- Another unique feature of water attributed to hydrogen bonding is its unusual density.

- Almost all substances have their highest density when they are in the solid state.

- But water is an exception as it has its maximum density at 4°C where it is liquid.

- Water in the solid state, ice has a lower density than the liquid water.

- At temperatures below 4°C, water forms a rigid open structure of hexagonal networks in which hydrogen bonding is maximized.

- Each oxygen atom is tetrahedrally surrounded by four other oxygen atoms, with hydrogen bonds bridging every pair of oxygen atoms.

- It is such arrangement which creates the open structure and makes it less dense.

- When water melts, the hydrogen bonding becomes more random, breaking down the lattice structure of ice for the water molecules to pack.

- More closely for the density to increase.

- The fact that ice is less dense than water has profound ecological significance.

- Water in rivers, lakes, oceans freeze from the surface to the bottom.

- The frozen water (ice) at the top insulates the water below from freezing, so that complete solidification of water bodies is prevented, and this is beneficial to aquatic flora and fauna in temperate regions, but not in our tropical climate.

- It is not advisable to keep glass bottles fully filled in fridge: why?

- Hydrogen bonding is also responsible for the viscosity of some molecules.

- The ability of a liquid to flow is called its viscosity.

- The higher the viscosity, the more the resistance to How, or the slower the flow.

- Viscosity is measured by observing how long it takes a given volume of liquid to flow through a narrow tube in a viscometer.

- Viscosity of a liquid arises from the intermolecular forces.

- The stronger the forces hindering the motion of the molecules, the greater the viscosity.

- Hydrogen bonding is particularly important in this respect because it can link neighbouring molecules together so strongly.

- This explains why water has greater viscosity than benzene:

- The latter has no hydrogen bonding phosphoric acid and glycerol are even more viscous than water at room temperature because of the numerous hydrogen bonds present.

- Engine oil is more viscous than water, but being a hydrocarbon, there is no hydrogen bonding.

- Two factors are responsible for the high viscosity of engine oil.

- Being of high molecular sizes, the van der Waals forces between the molecules are quite high.

- Then, the long chain-like molecules make the molecules get entangled with each other.

- Viscosities decrease with temperature.

- With the increasing temperature, the greater average kinetic energy of the molecules enables them overcome the attractive forces, allowing ease of flow.

- Another effect produced by hydrogen bonding is surface tension of liquids.
- Every molecule is attracted to the other molecules that surround it.
- Molecules in the bulk of the liquid are attracted on all sides and are not pulled in any one direction.

- On the other hand, the surface molecules have no molecules above them.

- Therefore, the surface molecules are drawn inward by the force of attraction called surface tension.

- Because of this attractive force on the surface, the liquid takes the shape that gives it as little surface area as possible.

- Surface tension produces the ‘skin effect’ on the surface of liquids.

- This is what allows a blade to float on water.

- Certain insects can also skate on water surface without sinking, even though they are denser than water.

- Mosquito larvae can be suspended in slow moving or stagnant water bodies for the same reason.

- Rain falls in the form of droplets because in that shape, the water has the least surface area.

- Glass containers can be filled above the brim with water without overflowing.

- Intermolecular forces are also responsible for capillary action, which is the rise of liquids up very narrow tubes.

- The liquid rises as a result of the attractive forces between the liquid and the inner surface of the tube.

- Such forces of attraction between the liquid molecules and the surface of the tube are the adhesive forces.

- Those forces that bind molecules together in the bulk, of a liquid are the forces of cohesion.

- The height a liquid rises is proportional to the surface tension (**γ**) and inversely proportional to the density of the liquid.

- $ph= 2 \boldsymbol{\gamma} /pgr$

- where *h is* height, *p* is density of liquid, g is the force of gravity and r is the radius of the tube.

- Whether the meniscus of a liquid will curve upward or downward in a capillary depends on the magnitude of the forces of adhesion and cohesion

- The meniscus of water in a tube is upward as the forces that make the water molecules adhere to the glass surface are stronger than the forces of cohesion between the water molecules.

- The meniscus of mercury is the tube, showing the forces of cohesion in the mercury is higher than the forces of to the glass.

- By the meniscus curving downward, the mercury tends to reduce its contact with the glass surface.

- Hydrogen bonding plays a major role in determining the molecular conformation of proteins and nucleic acids.

- The secondary structures or proteins are maintained by hydrogen helical structure of DNA is also possible because of the hydrogen bonding.

- Due to the weak nature of these hydrogen bonds, the two strands of DNA can unwind:

- this is an important step in the transfer of genetic information through replication of DNA.

- The higher degree of hydrogen bonding amongst the straight chain polymers of glucose in cellulose makes this polysaccharide more rigid than starch or glycogen which has lower degree of intermolecular hydrogen bonding.

- Bonding plays a major role in determining the molecular conformation of and nucleic acids.

- The secondary structures of proteins are maintained by hydrogen Examples of such secondary structure are α—helix and β-pleated sheets.

- The double structure helical structure of DNA is also possible because of hydrogen bonding.

- Due to the weak nature of these hydrogen bonds, the two strands of DNA can unwind:

- This is an important step in the transfer of genetic information through replication of DNA.

- The higher degree of hydrogen bonding amongst the straight chain polymers of glucose in cellulose makes this polysaccharide more rigid than starch or glycogen which has lower degree of intermolecular hydrogen bonding.