Chemical Bonding

→ **Ionic Bonding**

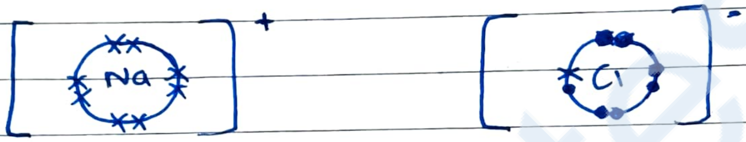
→ A bond between metal and non-metal.

→ Metals lose electrons and non-metals gain electrons.

→ This results in a very strong ~~force of attraction~~ electrostatic force of attraction between the metal cation and non metal anion.

→ **Dot and Cross diagram**

→ Example: NaCl



→ **Properties of ionic compounds**

→ very strong F.O.A

→ Solid at room temperature

→ High melting and boiling points

→ generally soluble in water, water molecules form bonds with ions replacing bonds within the lattice, thus ions go into the solution.

→ very good conductor of electricity (when in liquid/aqueous/molten state) it does not conduct when solid.

→ **Covalent Bonding**

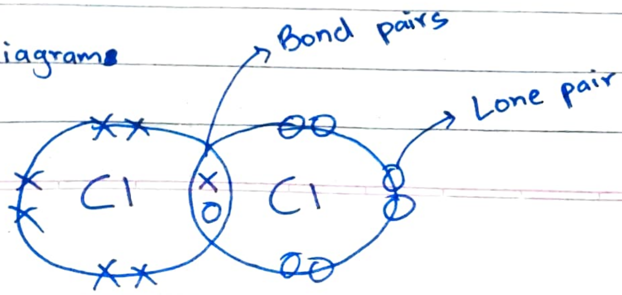
→ When two non-metals react they both need to gain electrons, so they do this by sharing electrons.

→ single bond is formed (Cl−Cl) when 1 pair of e^- is shared

→ double bond is formed (O=O) when 2 pairs of e^- are shared

→ triple bond is formed (N≡N) when 3 pairs of e^- are shared.

→ **Dot and Cross diagrams**



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⇒ Electronegativity and Bond polarity

→ Electronegativity is a measure for attraction of electrons, the higher the electronegativity the higher the attraction.

→ Electronegativity is the extent to which an atom covalently bonded to another atom attracts the bonded pair of electrons.

→ Example: HCl (Chlorine is more electronegative so it attracts the bonded pair of electron

H – Cl → it is called a dipole

Δ+ve ← δ^+ δ^- → Δ-ve

→ it is also called a polar bond (because it forms a dipole and the electronegativity of both elements/atoms is different.

→ Example Cl – Cl this is a non-polar bond because electroneg is same of both atoms. → C – H

→ Electronegativity increases across a period

→ Electronegativity decreases down a group

⇒ Intermolecular forces

→ Van Der Waals' force

  → Instantaneous dipole - induced dipole forces (temporary dipole)

  → Permanent dipole - dipole forces

→ Hydrogen Bonds.

⇒ Instantaneous dipole - induced dipole (id-id)

→ it exists between non-polar molecules

→ weakest type of intermolecular force

→ larger the molecule, stronger the id-id force.

→ more the electrons, stronger the id-id force.

⇒ Permanent dipole - dipole force (pd-pd)

→ it exists between polar molecules

→ stronger than id-id forces.

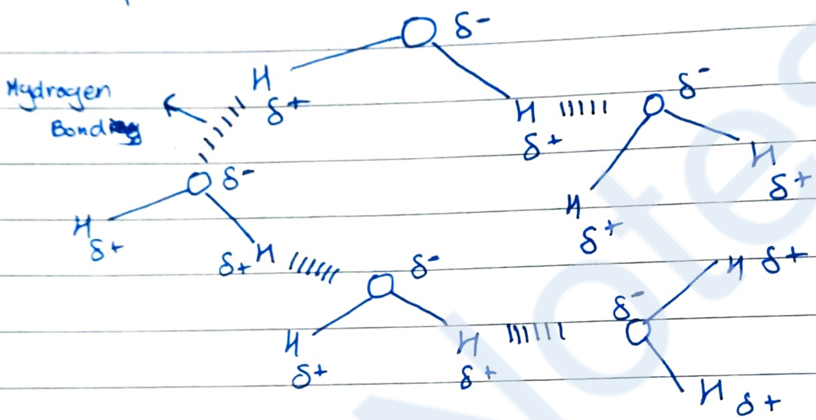
→ larger the molecule, stronger the pd-pd force

→ more the electrons, stronger the pd-pd force.

⇒ Hydrogen Bonding

→ strongest type of intermolecular force
→ Hydrogen has to be present for Hydrogen bonding to take place
→ it can only take place with highly polar compounds.
→ highly polar compounds form when hydrogen bonds with highly electronegative atom.
→ highly electronegative atoms → Fluorine, Oxygen, Nitrogen
HF   -OH   -NH

→ Example : Water



→ Properties of covalent compounds

→ Covalent bonds result in the formation of simple molecular structure
→ these molecules are held together by intermolecular forces
→ molecules that are held together by Van Der Waals' forces are usually gases and have low M.P and low B.P.
→ molecules that are held together by Hydrogen bonds have high M.P and high B.P.
→ LIKE DISSOLVES LIKE
  → molecules held by id-id forces (Non-polar) are insoluble in water (polar and hydrogen bond)
  → molecules held by pd-pd forces (polar) are weakly soluble in water (polar and hydrogen bond) and weakly soluble in non-polar solvents and highly soluble in polar solvents.
  → molecules held by hydrogen bonds are highly soluble in water
→ Overview - id-id dissolves in id-id
pd-pd dissolves in pd-pd
Hb-Hb dissolves in Hb-Hb

⇒ Properties of covalent compounds [continued]

→ simple molecular compounds do not conduct electricity because they have neither mobile ions nor mobile electrons.

→ water has a high surface tension compared to other liquids because of hydrogen bonds.

→ Ice is less dense than water because it has larger hydrogen bonds which occupies more space, which means increase in volume which means decrease in density.

→ H_2S should be stronger than H_2O as it has more electrons it is a larger molecule, but it has pd-pd force while H_2O has hydrogen bonds ∴ H_2O is stronger.

⇒ Shapes of molecules.

→ molecules have shapes due to electron repulsion

→ repulsion between 2 lone pairs is the highest

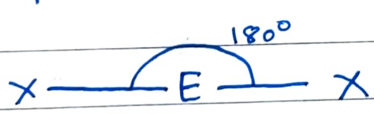
→ repulsion between lone pair and bond pair is smaller

→ repulsion between 2 bond pairs is the smallest.

LpLp > LpBp > BpBp

⇒ VSEPR (Valence Shell Electron Pair Repulsion)

1. 2 bond pairs 0 Lone pairs

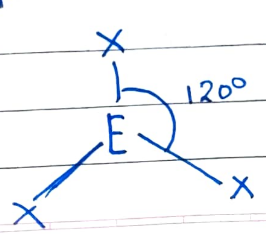


Example: $BeCl_2$

Name: Linear

Bond angle: 180°

2. 3 bond pairs 0 Lone pairs

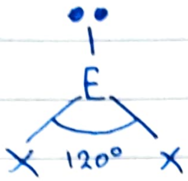


Example: BF_3

Name: Triagonal Planar

Bond angle: 120°

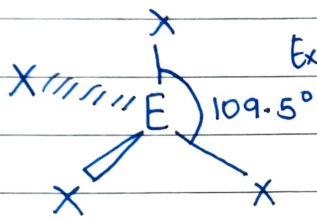
3. 2 bond pairs 1 Lone pair



Name : Bent or Angular

Bond angle : 120°

4. 4 bond pairs 0 Lone pairs

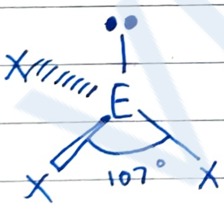


Ex. CH_4

Name : Tetrahedral

Bond angle : 109.5°

5. 3 bond pairs 1 Lone pair

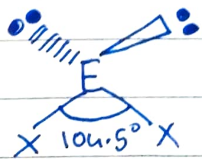


Ex. NH_3

Name : Triagonal Pyramidal

Bond angle : 107°

6. 2 bond pairs 2 Lone pairs

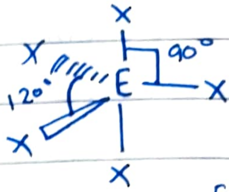


Ex. H_2O

Name : Bent or Angular

Bond angle : 104.5°

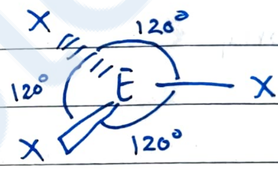
7. 5 bond pairs 0 Lone pairs



Ex. PCl_5

let's break this into 2

vertical



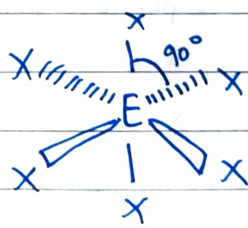
horizontal

This is triagonal planar.

∴ Name : Triagonal Bipyramidal

Bond angle : 90° and 120°

8. 6 Bond pairs 0 Lone pairs



Name : Octahedral

Bond angle : 90°

Example : SF_6

⇒ Hybridisation

→ Carbon = $1s^2 \; 2s^2 \; 2p^2$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow \,|\, \uparrow \,|\, \;}$ → Normal state

→ Carbon = $1s^2 \; 2s^1 \; 2p^3$

$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow \,|\, \uparrow \,|\, \uparrow}$ → Excited state

Carbon forms this excited state so that it can form 4 bonds and in this it will become more stable.

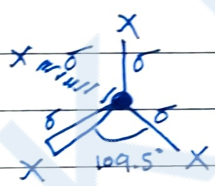
→ Hybridisation is when S-orbital and P-orbital merge.

→ Sp^3 $\boxed{\; | \; | \; | \;}$
s p p p

→ Sp^2 $\boxed{\; | \; | \;}$ $\boxed{\;}$
s p p p

→ Sp $\boxed{\; | \;}$ $\boxed{\; | \;}$ → Linear shape.
S p p p

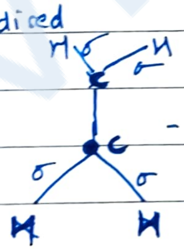
→ Sp^3 hybridised $\boxed{\uparrow\downarrow \,|\, \uparrow\downarrow \,|\, \uparrow\downarrow \,|\, \uparrow\downarrow}$
Sp^3 Sp^3 Sp^3 Sp^3



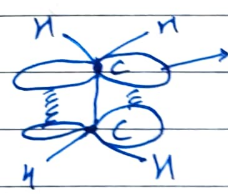
ex. CH_4 – Tetrahedral

all bonds are σ (sigma) bonds

→ Sp^2 hybridised



– Triagonal Planar



→ π (pie) bond between 2 p-orbitals.

both p-orbitals of each carbon overlap, to form a π bond

both carbons have formed σ bonds with H but both carbons have one p-orbital empty

$\boxed{\uparrow\downarrow}$
P

$\boxed{\uparrow\downarrow \,|\, \uparrow\downarrow \,|\, \uparrow\downarrow}$
Sp^2 Sp^2 Sp^2

$\boxed{\uparrow}$
p

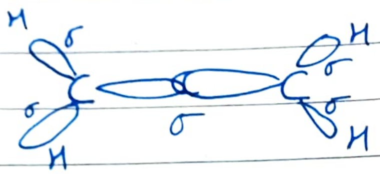
$\boxed{\uparrow\downarrow \,|\, \uparrow\downarrow \,|\, \uparrow\downarrow}$
Sp^2 Sp^2 Sp^2

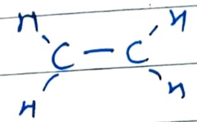
☆ lonely p-orbitals form π bonds.

⇒ σ and π bonds.

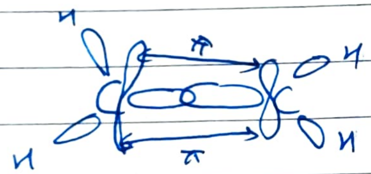
→ σ bonds

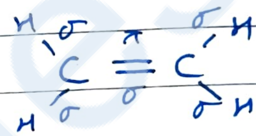
- → all single bonds have σ bonds
- → all bonds in sp^3 hybridisation are σ bonds. 4σ bonds
- → sp^2 hybridisation 3σ 1π



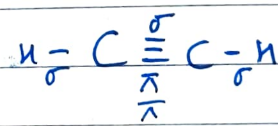


but now one p-orbital is unfilled for both carbon atoms.





→ sp hybridisation 2σ 2π



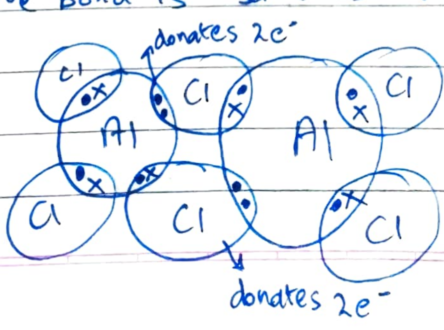
⇒ Bond energy

→ it is the energy needed to break one mole of a bond in gaseous state.

→ as bond length increases bond energy decreases.

⇒ Coordinate Bonding

→ sometimes one atom provides both the electrons for formation of a covalent bond, it is also called a dative bond. The strength of a dative bond is same as that of covalent bond.



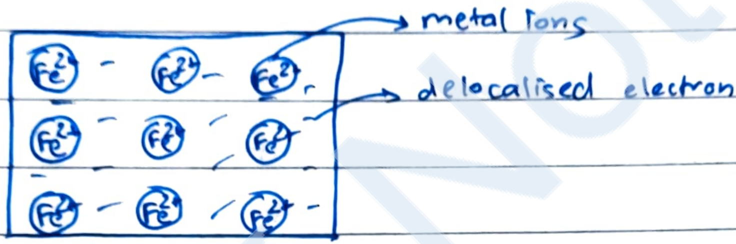
Al_2Cl_6 → Dimer

$AlCl_3$ → present at high temp.

⇒ Metallic Bonding

→ It consists of metal ions and delocalised sea of electrons.

→ there is a f.o.a between metal ions and delocalised sea of e^-.



⇒ Properties of metals

→ high M.P and B.P

→ malleable and ductile

→ good conductors of electricity.

→ insoluble. But some react with water

→ good conductors of heat