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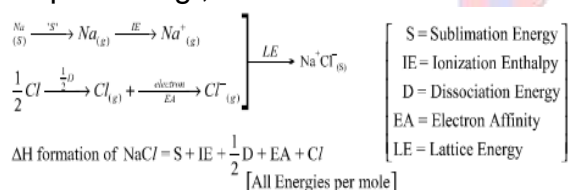
Chemical Bonding

- The interaction between two atoms which holds them together within a molecule or ions is known as **chemical bond**.
- The elements with one, two, three, four, five, six or seven electrons in outer shell, use these electrons to complete octet. The electrons which take part in two or more atoms to complete octet is known as **electrovalency**.
- Lewis symbols or electron dot symbols** involve the presentation of valence electrons (outer electrons) in an atom
 $\dot{\text{N}}\text{a}, \text{Mg}\cdot, \ddot{\text{A}}\text{l}\cdot, \ddot{\text{S}}\text{i}\cdot$ etc.
- Electrovalent bond or ionic Bond:** The bond (chemical interaction) between two atoms formed by complete transference of electron from valence shell (outer shell) of an atom to another to complete octet (noble gas configuration) [$2e^-$ in H, Li] is known as **ionic bond**.
- This ionic bond is favored by low ionization enthalpy of metal high electron gain enthalpy of non-metal atom and in the resulting ionic compound more lattice energy.
- Characteristics of ionic compound:** They are solids, a definite arrangement/pattern of ion (to give crystalline solids), high MP and BP, conductors in fused state and in aqueous medium, soluble in H_2O [Hydration].
- Octate rule:** During a chemical reaction the atoms tend to adjust their electronic arrangement in such a way that they achieve 8 e^- in their outermost electron.
- Chemical bond:** the chemical force which keeps the atoms in any molecule together is called a chemical bond.
- Ionic bond:** An ionic bond is formed by the complete transfer of one or more electrons from the atom of a metal to an atom of non-metal.
- Lattice enthalpy:** It is a measure of the strength of the forces between the ions in an ionic solid. The greater the **lattice enthalpy**, the stronger is the forces.

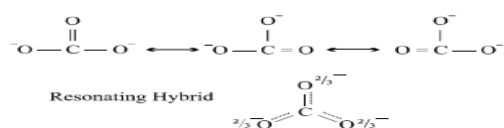
$$\begin{array}{ccc} \text{e.g., Mg} \longrightarrow \text{Mg}^{2+} + 2e^- & \text{O} + 2e^- \longrightarrow \text{O}^{2-} \\ (2, 8, 2) & (2, 6) \\ \text{Mg}^{2+} + \text{O}^{2-} \longrightarrow \text{MgO} \text{ Ionic compound [A crystalline lattice].} \end{array}$$
- Electro valency:** The number of electrons lost or gain by an atom of an element is called as electrovalency.
- Covalency:** The number of electrons which an atom contributes towards mutual sharing during the formation of a chemical bond.
- Single covalent bond:** A covalent bond formed by the mutual sharing of one pair of electrons is called a single covalent bond. It is represented by a small line (–) between the two atoms.
- Double covalent bond:** A covalent bond formed by the mutual sharing of two pair of electrons is called a double covalent bond. It is represented by two small horizontal lines (=) between the two atoms. E.g. $\text{O}=\text{O}$, $\text{O}=\text{C}=\text{O}$ etc.
- Triple covalent bond:** A covalent bond formed by the mutual sharing of three pair of electrons is called a triple covalent bond. It is represented by three small horizontal lines (\equiv) between the two atoms. E.g. $\text{N}\equiv\text{N}$, $\text{H}-\text{C}\equiv\text{C}-\text{H}$ etc.
- Polar covalent bond:** The bond between two unlike atoms which differ in their affinities for electrons is said to be polar covalent bond. E.g. H-Cl
- Coordinate bond:** The bond formed by the sided sharing of electrons take place is called a coordinate bond. It is represented by an arrow (\rightarrow) pointing towards the acceptor

atom. E.g. $\text{H}_3\text{N} \rightarrow \text{BF}_3$

- **Bond length:** Bond length is defined as the equilibrium distance between the nuclei of two bonded atoms in a molecule.
- **Bond angle:** It is defined as the angle between the orbitals containing bonding electron pairs around the central atom in a molecule/complex.
- **Bond enthalpy:** It is defined as the amount of energy required to break one mole of bonds of a particular type between two atoms in a gaseous state.
- **Bond order:** In the Lewis description of covalent bond, the Bond Order is given by the number of bonds between the two atoms in a molecule.
- **Born Haber Cycle:** For formation of ionic compound e.g., Na^+Cl^- .



- **Resonance:** (a) The delocalization of $2e^-$ in a molecule/ion which results in observed bond length, bond order, bond energy different from normal covalent bond. Data is known as resonance. Various resonating structures have nearly same energy and interconvertible to each other.
- (b) It gives stability to the molecule/ion. Atom does not shift their position in any of the resonating structure. The structure which is near to all resonating structure and nearly explains the property of that molecule/ion is known as **resonating hybrid**.



- **Dipole moment :** Dipole moment (μ) = Charge (Q) \times distance of separation (r)

- **Fajan's Rule:** Polarizability and polarizing power. The power of an ion to distort the other ion is polarizing power and the tendency of an ion to get distorted is known as polarizability. Factors affecting polarizing power and polarizability.

(a) High charge and small size of C^+ .

(b) High charge and large size of A^- .

- **Hybridization:** (a) Phenomenon of intermixing of atomic orbitals with slightly different energies to form new orbitals of equal energy and identical shape. The new orbitals are known as hybrid orbitals.

(b) The number of hybrid orbitals is equal to the number of orbitals involved in hybridization.

(c) As the 's' character in hybrid orbitals increases, electronegativity and size of hybrid orbitals decrease.

(d) Hybrid orbitals can have one unpaired electron as well as two unpaired electrons.

- **VSEPR (Valence shell electron pair repulsion theory):**

(a) There are three types of repulsion in a covalent molecule: $lp - lp > lp - bp > bp - bp$.

(b) These repulsions depend upon the difference in electronegativity between A and B.

(c) The hybridization of the central atom decides the geometry, and VSEPR theory decides the shape of the molecule.

- **Formal charge:** The charge of an atom in a molecule/ion.

$$\text{FC} = [\text{Total no. of valence in free atom}] - [\text{Total no. of non bonding electrons}] - \frac{1}{2} [\text{Total no. of shared electrons}]$$

- **The valence bond approach :** (a) The two atomic orbitals, each with one electron, overlap to give a region of maximum electron density common to both atoms, known as a single covalent bond.



(b) The strength of covalent bond is proportional to extent of overlapping between the atomic orbitals of valence shell.

- **Sigma and Pi π bond:** (a) The bond formed by overlap of two atomic orbitals along the internuclear axis of two atoms is Sigma bonds.

Summary of the chapter

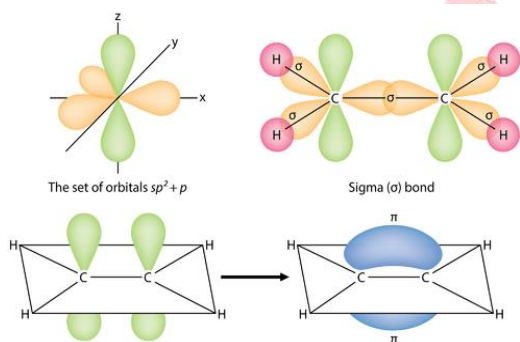
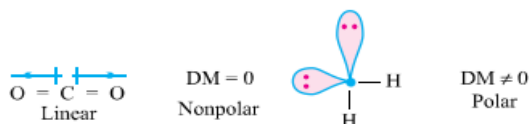


Fig.4.1: Formation of Sigma and Pi π bond

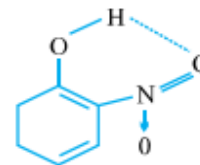
- **Dipole moment:** (a) For polar covalent molecules (atoms with difference in electronegativity) the product of charge separation and distance b/w charges is known as dipole moment.
- (b) Being vector, if net resultant of all vectors is zero the molecule has zero DM and known as non polar otherwise.



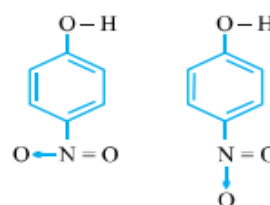
- **Hydrogen bonding:** The dipole interaction b/w molecules when H is bounded with highly electronegative atoms (F, O, N only).

(a) **Intermolecular HB:** When hydrogen bonding is in between the same molecule. Hence molecules are independent and with less MP, BP. Due to within hydrogen bonding notable to

make hydrogen bonding with H₂O therefore less soluble in water.



- (b) **Intermolecular hydrogen bonding:** When hydrogen bonding in between the different molecules therefore close packing of molecules therefore more MP and BP and more soluble in water.



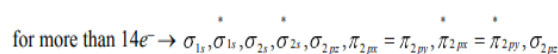
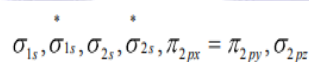
- **Molecular orbital theory :** (a) The overlap of atomic orbitals of same symmetry of two homonuclear atoms to give addition or subtraction of wave functions and form bonding MO and antibonding MO respectively is known as MO theory.

(b) The e^- is filled in molecule increasing energy order of MO.

(c) **Bond order:** Bond order of molecule/ion.

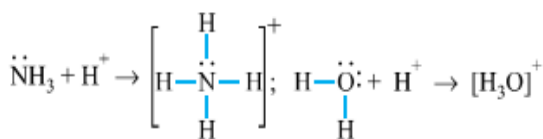
$$= \frac{\text{No. of electron in BMO} - \text{No. of electrons in ABMO}}{2}$$

- (d) More is bond order more is bond energy lesser is the bond length. Bond order zero means no. possibility of that molecule. (e) Increasing order of energy of MO for upto $14e^-$.





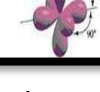


(e) This theory decides the magnetic behaviour also. [Equal energy orbitals]

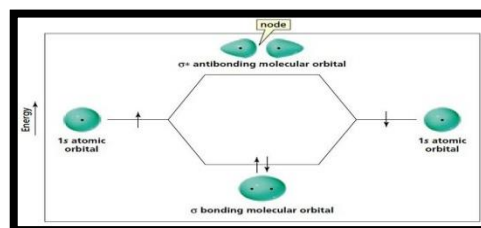
- **Coordinate covalent bond:** The sigma bond formed by donation of lp into vacant by drized orbital of other atom (acceptation atom) is known as **coordinate covalent bond** or donor acceptor or **daive bond**.



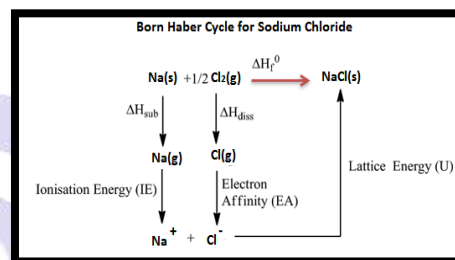
- **Hybridisation** (or hybridization) is a process of mathematically combining two or more atomic orbitals from the same atom to form an entirely new orbital different from its components and hence being called as a hybrid orbital.

Valence electron pair geometry	# of orbitals	Hybrid orbitals	Electron density diagram	Examples
Linear	2	sp		BF ₂ HgCl ₂ CO ₂
Trigonal planar	3	sp ²		BF ₃ SO ₃
Tetrahedral	4	sp ³		CH ₄ H ₂ O NH ₄ ⁺
Trigonal bipyramidal	5	sp ³ d		PF ₅ SF ₄ BrF ₃
Octahedral	6	sp ³ d ²		SF ₆ XeF ₄ PF ₆ ⁻

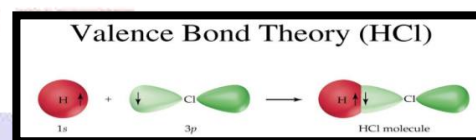
- Valence shell electron pair repulsion **theory**, or **VSEPR theory**, is a **model** used in chemistry to predict the geometry of individual molecules from the of electron pairs surrounding their central atoms.
- **Molecular orbital theory** (MO theory or MOT): It is a method for describing the electronic structure of molecules using quantum mechanics. It was proposed early in the 20th century. **Molecular orbital theory** and valence bond **theory** are the foundational **theories** of quantum chemistry.



- **The Born-Haber cycle:** It is a classic approach to measure the reaction energies. The cycle is mostly concerned with the formation of an ionic solid from the metals (Group I or Group II) when reacted with a halogen or a non-metallic element like oxygen.



- **Valence bond theory:** It is a **theory** which describes chemical bonding. **VBT** states that the overlap of incompletely filled atomic orbitals leads to the formation of a chemical bond between two atoms. The unpaired



electrons are shared and a hybrid orbital is formed.

- **Bond order** = [(Bonding molecules' number of electrons) – (Antibonding molecules' number of electrons)]/2. I.e. **Bond Order** = ½ [Nb – Na] Where, Nb is that the number of **bonding** electrons.
- **Lewis Structure:** It is a very simplified representation of the valence shell electrons in a molecule. It is used to show how the electrons are arranged around individual atoms in a molecule. Electrons are shown as "dots" or for bonding electrons as a line between the two atoms

Check Yourself

1. C-O bond length is minimum in
(A) CO_2 (B) CO_3^{2-}
(C) HCOO^- (D) CO
2. Molecules are held together in a crystal by
(A) Hydrogen bond
(B) Electrostatic attraction
(C) Van der Waal's attraction
(D) Dipole-dipole attraction
3. Sp^3d^2 hybridization is present in $[\text{Co}(\text{NH}_3)_6^{3+}]$, find its geometry
(A) Octahedral geometry
(B) Square planar geometry
(C) Tetragonal geometry
(D) Tetrahedral geometry
4. Find the molecule with the maximum dipole moment
(A) CH_4 (B) NH_3 (C) CO_2 (D) NF_3
5. MX_6 is a molecule with octahedral geometry. How many X – M – X bonds are at 180° ?
(A) Four (B) Two
(C) Three (D) Six

Test Yourself

Q. Draw all the atomic models given in your textbook and compare which one is the correct explanation for the atomic structure.

Also discuss the Drawbacks of some atomic models.

Stretch Yourself

1. Explain the non linear shape of H_2S and non planar shape of PCl_3 using valence shell electron pair repulsion theory.
2. Using molecular orbital theory, compare the bond energy and magnetic character of O^{2+} and O^{2-} species.
3. Explain the shape of BrF_5 .
4. Explain why PCl_5 is trigonal bipyramidal whereas IF_5 is square pyramidal.
5. In both water and dimethyl ether ($\text{CH}_3\text{—}\ddot{\text{O}}\text{—}\text{CH}_3$), oxygen atom is central atom, and has the same hybridization, yet they have different bond angles.



Answers

Check Yourself

Answer: 1(D); 2(C); 3(A); 4(B); 5(C)

Stretch Yourself

- In both H_2S and PCl_3 , lone pairs are present along with bond pairs around the central atom. According to VSEPR, theory $\text{L.P-L.P} > \text{L.P-B.P} > \text{B.P.-B.P}$. Therefore, shapes are different.
- In O_2^+ , $\text{B.O} = (10-5)/2=2.5$ In O_2^- , $\text{B.O} = (10-7)/2=1.5$ Bond dissociation energy is directly proportional to bond order and paramagnetic character depends on unpaired electrons. O_2^+ has high dissociation energy and both are paramagnetic.
- BrF_5 : Br is surrounded by 5 bonded pairs and one lone pair. Shape is square pyramidal.



Structure of BrF_5

- In PCl_5 , P is surrounded by 5 bond pairs and no lone pairs, whereas in IF_5 , iodine atom is surrounded by 5 bond pairs and one lone pair, therefore the shape is different.
- Dimethyl ether will have larger bond angle. There will be more repulsion between bond pairs of $-\text{CH}_3$ groups attached in ether than between bond pairs of hydrogen atoms attached to oxygen in water. The carbon of $-\text{CH}_3$ in ether is attached to three hydrogen atoms through bond and electron pair of these bonds add to the electronic charge density on carbon atom. Hence, repulsion between two $-\text{CH}_3$ groups will be more than that between two hydrogen atoms.