

Types:

- (i). **Covalent Bond:** A chemical bond formed between two atoms by mutual sharing of electrons between them to complete their octet.
(ii). **Ionic Bond:** A chemical bond formed by complete transference of electrons from one atom to another acquire the stable nearest noble gas configuration.

Energy required to completely separate one mole of a solid ionic compound into gaseous constituent ions.

Postulates:

- Shape of molecule depends upon the number of valence shell electron pairs around central atom.
- Pairs of electrons in the valence shell repel one another.
- These pairs of electrons tend to occupy such positions in space that minimize repulsion.
- The valence shell is taken as a sphere with electron pairs localising on spherical surface at maximum distance from one another.
- A multiple bond is treated as if it is a single electron pair
- and the two or three electron pairs of a multiple bond are treated as a single super pair.
- When one or more resonance structures can represent a molecule, VSEPR model is applicable:
- **Decreasing order of repulsive interaction:**
lp-lp > lp-s > s-s

Valence Bond Theory: Given by L. Pauling. It explains that a covalent bond is formed between two atoms by overlap of their half-filled valence orbitals, each of which contains one unpaired electron.

Orbital Overlap Concept: Formation of a covalent bond results by pairing of electrons in valence shell with opposite spins.

Types of Overlapping: (i) Sigma σ bond - end to end.
(ii) Pi (π) bond - axis remain parallel to each other.

Hybridisation: Process of intermixing of orbitals of different energies resulting in formation of new set of orbitals of equivalent energies and shape.

Types of Hybridisation: (i) sp (ii) sp² (iii) sp³

Bonding Molecular Orbitals: Addition of atomic orbitals.
Antibonding Molecular Orbitals: Subtraction of atomic orbitals.

- (i) **Bond Length:** Equilibrium distance between the nuclei of two bonded atoms in molecule.
(ii) **Bond Angle:** Angle between the orbitals containing bonding electron pairs around central atom in a molecule complex ion.
(iii) **Bond Enthalpy:** Amount of energy required to break one mole of bonds of particular type between two atoms.
(iv) **Bond Order:** Number of bonds between the two atoms of a molecule
(v) **Resonance Structures:** Are a set of two or more Lewis structures that collectively describe the electronic bonding a single polyatomic species.
(vi) **Dipole Moment:** Product of the magnitude of the charge and distance between centres of positive and negative

Chemical Bond: Attractive forces which hold the various chemical constituents together in different chemical species.

Lattice Enthalpy

Molecular orbital theory

Chemical Bonding and Molecular Structure

Valence Shell Electron Pair Repulsion (VSEPR)

Chemical Bonding

Hybridization

Bond Parameters

Postulates:

- Electrons in a molecule are present in various molecular orbitals as electrons are present in atomic orbitals.
 - Atomic orbitals of comparable energies and proper symmetry combine.
 - Atomic orbitals are monocentric while a molecular orbital is polycentric.
 - Number of molecular orbital formed is equal to number of combining molecular orbitals.
 - Bonding molecular orbitals have low energy and high stability
- Types of MO: σ (Sigma), π (Pi), δ (Delta)

Kossel Lewis approach to chemical bonding:

Lewis pictured the atom as a positively charged 'kernel' and the outer shell accommodates a maximum of eight electrons.

- Lewis postulated that atoms achieve the stable octet when linked by chemical bonds.
- Kossel gave following facts:
In the periodic table, highly electronegative halogens and highly electropositive alkali separated by noble gases.
Formation of a negative ion from a halogen atom and a positive ion from an alkali metal atom is associated with gain and loss of electron by respective atoms.
Negative and positive ions formed attain noble gas electronic configurations.
- Negative and positive ions are stabilized by electrostatic attraction.

Octet Rule: Atoms can combine either by transfer of valence electrons from one atom to another or by sharing of valence electrons to complete octet in their valence shells.

Lewis Dot Structure provides a picture of bonding in molecules and ions in terms of the shared pairs of electrons and the octet rule.

How To Write A Lewis Dot Structure:

- Step 1: Add the valence electrons of the combining atoms to obtain total number of electrons.
 - Step 2: For anions, each negative charge means addition of one electron. For cations, each positive charge means subtraction of one electron from total number of valence electrons.
 - Step 3: Write chemical symbols of combining atoms.
 - Step 4: Least electronegative atom occupies central position.
 - Step 5: After accounting for shared pairs of electrons remaining are either utilized for multiple bonding or remain as lone pairs.
- Formal Charge = (Total number of valence electrons in free atom) - (Total number of non-bonding electrons) - 1/2(Total number of bonding electrons)

Limitations Of Octet Rule:

- Shows three types of exceptions (ie) incomplete octet of central atom, odd-electron molecules and expanded octet.
- Does not account for the shape of molecules.
- Fails to explain stability of molecules.

Hydrogen Bond: Formed when the negative end of one molecule attracts the positive end of other.

Types:

- (i) **Intermolecular:** Between two different molecules of same or different compounds.
(ii) **Intramolecular:** H atom is between two highly electronegative atoms.

