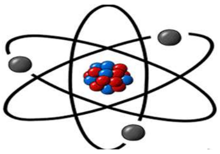
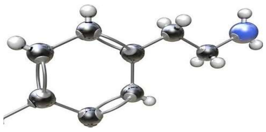
CHEMICAL BONDING & MOLECULAR STRUCTURE





NCERT | QUICK REVISION NOTES

1. CHEMICAL BONDING

Attractive force holding atoms/ions together.

Why bonds form?

to achieve greater stability and lower energy, by attaining a noble-gas configuration.

9. BOND PARAMETERS

Bond length

average distance between the nuclei of two bonded atoms.

C–C > C=C > C≡C

Bond enthalpy

energy required to break one mole of a particular bond

C–C < C=C < C≡C

Bond order

number of bonds between two atoms.

C–C < C=C < C≡C

Bond Angle

angle between two bonds formed around the central atom

Bond Strength

how strongly two atoms are held together by a chemical bond

C–C < C=C < C≡C

10. SHAPES OF MOLECULES

Molecule	Shape	Bond angle
$BeCl_2$	Linear	180°
BF_3	Trigonal planar	120°
CH_4	Tetrahedral	109.5°
NH_3	Trigonal pyramidal	107°
H_2O	Bent	104.5°
PCl_5	Trigonal bipyramidal	90°, 120°
SF_6	Octahedral	90°

3. IONIC BOND

Complete transfer of electrons

Example:

$Na \rightarrow Na^+ + e^-$

$Cl + e^- \rightarrow Cl^-$

$Na^+ + Cl^- \rightarrow NaCl$

high melting point, crystalline, conduct electricity in molten/aqueous state, electrostatic attraction.

2. OCTET RULE

Atoms tend to achieve 8 electrons in their valence shell.

Exceptions:

Incomplete octet: BF_3

Odd-electron molecules: NO, NO_2

Expanded octet: PCl_5, SF_6

6. KOSSEL-LEWIS APPROACH

explains chemical bonding on the basis of the electronic configuration of atoms by transfer of electrons.

Example: Na:2,8,1

$Na \rightarrow Na^+ + e^-$

Cl:2,8,7

$Cl + e^- \rightarrow Cl^-$

$Na^+ + Cl^- \rightarrow NaCl$

8. FORMAL CHARGE

$$FC = V - \left(L + \frac{B}{2}\right)$$

where:

- V = valence electrons
- L = non-bonding electrons
- B = bonding electrons.

4. COVALENT BOND

Mutual sharing of electron pairs.

- Single → 1 shared pair
- Double → 2 shared pairs
- Triple → 3 shared pairs

Example:

H–H , O=O , N≡N

low melting and boiling point, poor conductors

5. COORDINATE BOND

Both electrons of the shared pair are donated by one atom.

Example:

$NH_3 + H^+ \rightarrow NH_4^+$

Also called Dative Bond

7. LEWIS STRUCTURE

The valence electrons are represented by dots around the symbol of an element.

Example: Na· , :Cl:

11. VSEPR THEORY

Valence Shell Electron Pair Repulsion Theory

Electron pairs present around the central atom repel each other and arrange themselves as far apart as possible to minimise repulsion.

The repulsion follows:

LP–LP > LP–BP > BP–BP

12. IMPORTANT COMPARISON

Molecule	Lone pairs	Shape	Bond angle
CH_4	0	Tetrahedral	109.5°
NH_3	1	Trigonal pyramidal	107°
H_2O	2	Bent	104.5°

The bond angle decreases because lone-pair repulsion increases.

$CH_4 > NH_3 > H_2O$

13. KEYPOINTS

- More lone pairs generally mean smaller bond angles.
- Higher bond order means shorter and stronger bonds.
- Resonance increases stability.
- Dipole moment helps determine molecular polarity.
- Symmetrical molecules may have zero net dipole moment even when individual bonds are polar.

14. MOT

Atomic orbitals combine to form molecular orbitals.

Bonding Molecular Orbital

- Lower energy
- Increases stability

σ, π

Antibonding Molecular Orbital

- Higher energy
- Decreases stability

σ^*, π^*

Bond Order

$BO = \frac{Nb - Na}{2}$

If $BO > 0$

the molecule can be stable.

15. HYBRIDISATION

mixing of atomic orbitals to form equivalent hybrid orbitals.

Hybridisation	Orbitals mixed	Geometry	Example
sp	1s + 1p	Linear	$BeCl_2$
sp^2	1s + 2p	Trigonal planar	BF_3
sp^3	1s + 3p	Tetrahedral	CH_4
sp^3d	1s + 3p + 1d	Trigonal bipyramidal	PCl_5
sp^3d^2	1s + 3p + 2d	Octahedral	SF_6

16. VALENCE BOND THEORY

a covalent bond is formed by the overlap of atomic orbitals containing unpaired electrons.

The overlapping orbitals contain electrons with opposite spins.

The greater the overlap, the stronger the bond generally becomes.

Types of Orbital Overlap

1. Sigma (σ) bond

head-on overlap

2. Pi (π) bond

sidewise/lateral overlap

17. MAGNETIC BEHAVIOUR

whether a molecule is paramagnetic or diamagnetic.

Paramagnetic

A substance containing one or more unpaired electrons is paramagnetic.

Diamagnetic

A substance with all electrons paired is diamagnetic.

18. HYDROGEN BONDING

attractive interaction involving H attached to a highly electronegative atom like F, O, N

Examples:

HF, H_2O, NH_3, alcohols

Intermolecular H-bond

Between different molecules.

Intramolecular H-bond

Within the same molecule.

Important Comparison:

$H_2O > H_2S$

19. IMPORTANT EXAMPLES

H_2, O_2, F_2, CO, N_2, NH_3, H_2O etc