

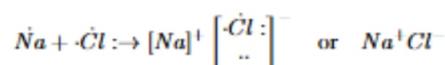
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Chemical Bonding

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1. Electrovalent Bond (Ionic Bond)

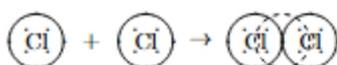
An ionic bond forms when electrons are completely transferred from a metal to a non-metal, forming oppositely charged ions held together by electrostatic attraction.



Examples: Na_2S , CaH_2 , AlF_3 , $MgCl_2$, $CaCl_2$, MgO .

2. Covalent Bond

Covalent bonds result from sharing of electron pairs between non-metal atoms to achieve stable configurations.



Electron configs: $(2,8,7) + (2,8,7) \rightarrow (2,8,8) + (2,8,8)$ or $Cl-Cl$

Examples: HCN , PCl_3 , PH_3 , H_2S , NH_3 .

3. Properties of Covalent Compounds

- Poor conductors of electricity.
- Insoluble in polar solvents; dissolve in non-polar solvents.
- Show isomerism due to directional bonding.
- Undergo slow molecular reactions.

4. Dipole Moment

$$\mu = q \times d$$

Where μ is in Debye: $1D = 3.33 \times 10^{-30}$ C-m.

Formula	Geometry	Dipole	Example
AX	Linear	Non-zero	HF, HCl
AX ₂	Linear/Bent	Zero/Non-zero	CO ₂ , H ₂ O
AX ₃	Planar/Pyramidal	Zero/Non-zero	BF ₃ , NH ₃
AX ₄	Tetrahedral/See-saw	Zero/Non-zero	CH ₄ , SF ₄
AX ₅	Trigonal bipyramidal/Sq. pyramidal	Zero/Non-zero	PCl ₅ , BrCl ₅
AX ₆	Octahedral/Distorted	Zero/Non-zero	SF ₆ , XeF ₆
AX ₇	Pentagonal bipyramidal	Zero	IF ₇

5. Fajan's Rule

1. Smaller cation: More polarizing.
2. Larger anion: More polarizable.
3. Higher charge: Greater covalent character.
4. Pseudo-noble gas config: More distortion.

6. Valence Bond Theory (VBT)

1. Bonds form via orbital overlap with opposite spins.
2. Greater overlap \Rightarrow stronger, shorter bond.

3. σ -bond: Head-on; π -bond: Sidewise overlap.

Sigma (σ) bond	Pi (π) bond
End-to-end overlap	Sidewise overlap
Stronger, more stable	Weaker, less stable
Exists independently	Accompanies σ bond
Symmetrical around axis	Electron cloud above/below axis

7. Hybridization

Mixing of similar-energy orbitals to form equivalent hybrid orbitals.

Key Points:

- Occurs during bonding only.
- Hybrid orbitals form σ -bonds.
- Count of hybrid orbitals = orbitals mixed.
- No π -bond orbitals are hybridized.

8. Resonance

When more than one Lewis structure is possible, actual structure is a resonance hybrid.

Resonance Energy:

$$E_{res} = E_{hybrid} - E_{most\ stable\ structure}$$

9. Bond Characteristics

Bond Length

- Larger atoms \Rightarrow longer bond.
- More bond order \Rightarrow shorter bond.
- Higher s-character \Rightarrow shorter bond.
- Polarity decreases bond length.

Bond Energy

Energy required to break 1 mol bond in gaseous state. Stronger bond \Rightarrow higher energy.

Bond Angle

- More lone pairs \Rightarrow angle reduces.
- Greater s-character \Rightarrow angle increases.

Type	sp^3	sp^2	sp
Angle	109.5°	120°	180°

Type	Lone Pairs	Hybrid	Angle	Expected	Actual
AX ₃	1	sp^2	< 120°	Trig. planar	Bent
AX ₄	2	sp^3	< 109.5°	Tetrahedral	Angular
AX ₅	1	sp^3d	< 120°	Trig. bi-pyramidal	See-saw
AX ₅	2	sp^3d	~ 90°	Trig. bi-pyramidal	T-shaped
AX ₅	3	sp^3d	180°	Trig. bi-pyramidal	Linear
AX ₆	1	sp^3d^2	< 90°	Octahedral	Sq. pyramidal
AX ₆	2	sp^3d^2	90°	Octahedral	Square planar

Examples: NH_3 , H_2O , SF_4 , ClF_3 , XeF_2 , XeF_4 , BrF_5 .

10. Molecular Orbital Theory (MOT)

1. AO's combine to form equal number of MO's.
2. Two MO types: Bonding and Antibonding.
3. Bonding MO < energy than AO; Antibonding MO > energy.
4. MO's filled per Aufbau principle.

11. Hydrogen Bonding

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Hydrogen bond is a dipole-dipole interaction between hydrogen and highly electronegative atoms (F, O, N).

Types:

- Intermolecular (e.g., H_2O)
- Intramolecular (e.g., salicylic acid)