CLASS 11 CHEMISTRY — FORMULA SHEET

Chapter 4 : Chemical Bonding and Molecular Structure (NCERT)

4.1 Lewis Structures, Octet Rule & Formal Charge

Concept	Formula / Rule
Octet rule	Atoms bond to attain 8 electrons in the valence shell (noble-gas configuration)
Formal charge (FC)	FC = V − L − ½B (V = valence e^- of free atom, L = lone-pair e^-, B = bonding/shared e^-)
Best Lewis structure	The one with formal charges closest to ZERO; negative FC on the more electronegative atom
Incomplete octet	LiCl, BeH_2, BCl_3, BF_3 (central atom has < 8 e^-)
Expanded octet	PF_5 (10 e^-), SF_6 (12 e^-), H_2SO_4, IF_7 (only 3rd period onward — d-orbitals)
Odd-electron molecules	NO (11 e^-), NO_2 (17 e^-) — octet rule fails
Resonance	Real structure = hybrid of all canonical forms; more resonance → more stability (O_3, CO_3^{2-}, benzene)

4.2 Ionic Bond, Lattice Enthalpy & Fajan's Rules

Concept	Rule / Statement
Ionic bond formation favoured by	Low ionization enthalpy of metal + high (negative) electron gain enthalpy of non-metal + high lattice enthalpy
Lattice enthalpy	Energy to completely separate 1 mole of ionic solid into gaseous ions; ↑ with charge, ↓ with ionic radius
Fajan's rules (covalent character ↑ if)	Cation: small & highly charged ; Anion: large & highly charged ; cation with pseudo noble-gas config (e.g. Cu^+)
Polarising power order	$Li^+ > Na^+ > K^+$ (smaller cation polarises more) ; covalent character: LiCl > NaCl > KCl

4.3 Bond Parameters & Dipole Moment

Property	Formula / Relation
Bond order (VBT)	= number of bonds between two atoms (C−C: 1, C=C: 2, C≡C: 3)
Bond order relations	Bond order ↑ → Bond enthalpy ↑ & Bond length ↓
Isoelectronic species	Same bond order: N_2, CO, NO^+, CN^- (all BO = 3)
Dipole moment (μ)	μ = q × d ; unit: Debye (D), 1 D = 3.336 × 10^{-30} C·m
% Ionic character	= (Observed μ ÷ Theoretical μ for 100% ionic) × 100
Zero dipole (symmetric)	CO_2, BF_3, CH_4, CCl_4, PCl_5, SF_6 — bond dipoles cancel
Non-zero dipole	H_2O (1.85 D), NH_3 (1.47 D) > NF_3 (0.23 D — lone pair opposes N−F dipoles)
Dipole order trick	cis-isomer has μ > 0; trans (symmetric) often μ = 0

4.4 VSEPR Theory — Geometry from BP + LP

BP + LP	BP	LP	Shape	Angle	Example
2	2	0	Linear	180°	$BeCl_2$, CO_2
3	3	0	Trigonal planar	120°	BF_3
3	2	1	Bent / V-shape	< 120°	SO_2, $SnCl_2$
4	4	0	Tetrahedral	109.5°	CH_4, NH_4^+
4	3	1	Trigonal pyramidal	107°	NH_3
4	2	2	Bent	104.5°	H_2O
5	5	0	Trigonal bipyramidal	120°, 90°	PCl_5
5	4	1	See-saw	—	SF_4
5	3	2	T-shape	—	ClF_3
5	2	3	Linear	180°	XeF_2, I_3^-
6	6	0	Octahedral	90°	SF_6
6	5	1	Square pyramidal	—	BrF_5
6	4	2	Square planar	90°	XeF_4

VSEPR Rule	Statement
Repulsion order	Lone pair–Lone pair > Lone pair–Bond pair > Bond pair–Bond pair
Effect of lone pairs	Each lone pair squeezes bond angle: CH_4 (109.5°) → NH_3 (107°) → H_2O (104.5°)

4.5 Valence Bond Theory & Hybridisation

Concept	Formula / Rule
Hybridisation shortcut	H = ½ (V + M − C + A) ; V = valence e^- of central atom, M = monovalent atoms, C = cationic charge, A = anionic charge
H value → type	2 → sp ; 3 → sp^2 ; 4 → sp^3 ; 5 → sp^3d ; 6 → sp^3d^2
Alternative count	Hybridisation = σ-bonds + lone pairs on central atom
Sigma vs pi	σ: head-on overlap (stronger, free rotation) ; π: sideways overlap (weaker); double bond = 1σ + 1π, triple = 1σ + 2π
% s-character & angle	sp (50% s, 180°) > sp^2 (33%, 120°) > sp^3 (25%, 109.5°); more s-character → shorter, stronger bond
Bond strength of overlap	s–s < s–p < p–p (for σ overlap of same period)

Hybridisation	Geometry	Bond Angle	Examples
sp	Linear	180°	$BeCl_2$, CO_2, C_2H_2
sp^2	Trigonal planar	120°	BF_3, C_2H_4, SO_3
sp^3	Tetrahedral	109.5°	CH_4, CCl_4, NH_4^+
sp^3d	Trigonal bipyramidal	120° & 90°	PCl_5
sp^3d^2	Octahedral	90°	SF_6
d^2sp^3 / sp^3d^2	Octahedral	90°	$[Co(NH_3)_6]^{3+}$ (Class 12 preview)

4.6 Molecular Orbital Theory (MOT)

Concept	Formula / Rule
Bond order (MOT)	BO = ½ (N_b − N_a) ; N_b = bonding e^-, N_a = antibonding e^-
Stability	BO > 0 → molecule exists ; BO = 0 → does NOT exist (He_2, Be_2)
Magnetism	Unpaired e^- → paramagnetic ; all paired → diamagnetic
Energy order (up to N_2, Z ≤ 14)	$\sigma1s < \sigma^*1s < \sigma2s < \sigma^*2s < (\pi2p_x = \pi2p_y) < \sigma2p_z < (\pi^*2p_x = \pi^*2p_y) < \sigma^*2p_z$
Energy order (O_2, F_2)	$\sigma2p_z$ comes BEFORE $\pi2p_x = \pi2p_y$
Adding to antibonding MO	Removes ½ from BO → O_2^+ (BO 2.5) > O_2 (2) > O_2^- (1.5) > O_2^{2-} (1)

Species	Total e^-	Bond Order	Magnetism
H_2	2	1	Diamagnetic
He_2	4	0 (does not exist)	—
Li_2	6	1	Diamagnetic
B_2	10	1	Paramagnetic (2 unpaired)
C_2	12	2	Diamagnetic
N_2	14	3	Diamagnetic
O_2	16	2	Paramagnetic (2 unpaired)
F_2	18	1	Diamagnetic
Ne_2	20	0 (does not exist)	—

4.7 Hydrogen Bonding

Concept	Rule / Example
Condition	H attached to highly electronegative F, O, or N
Strength order	H…F > H…O > H…N (but weaker than covalent bond, about 10-40 kJ/mol)
Intermolecular H-bond	Between molecules: H_2O, HF, alcohols → high b.p., water's anomalous density
Intramolecular H-bond	Within one molecule: o-nitrophenol, salicylaldehyde → LOWER b.p. than para-isomer
Why ice floats	Open cage structure due to H-bonding → ice less dense than liquid water

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FC in seconds	FC = V − L − B/2 → 'Valence minus Lone minus half-Bonded'.
Hybridisation trick	H = ½(V + M − C + A); SF_4: ½(6+4) = 5 → sp^3d ; NH_4^+: ½(5+4−1) = 4 → sp^3.
BO ladder for O_2 family	O_2^+ 2.5 > O_2 2 > O_2^- 1.5 > O_2^{2-} 1 — each added e^- goes to π* and cuts 0.5.
Paramagnetic pair	'B_2 and O_2 are the odd ones out' — both paramagnetic with 2 unpaired electrons.
Angle squeeze	109.5° → 107° → 104.5° : each lone pair eats 2.5° (CH_4 → NH_3 → H_2O).
NH_3 vs NF_3 dipole	NH_3 (1.47 D) >> NF_3 (0.23 D): lone pair adds in NH_3 but opposes in NF_3.
Zero-dipole list	'CO_2, BF_3, CH_4, CCl_4, PCl_5, SF_6' — perfectly symmetric = μ = 0.

Isoelectronic BO = 3 club	N_2, CO, NO^+, CN^- — all 14 e^-, all bond order 3.