3.1 Intermolecular Forces (IMFs)

IMF Types

IMF	Between	Key sign
LDF	all particles	e^- cloud size
Dipole–dipole	polar molecules	net dipole
H-bond	H on N/O/F + N/O/F	N–H, O–H, F–H
Ion–dipole	ion + polar	salt in water

London dispersion (LDF): temporary dipole induces a dipole in a neighbor. In **every** molecule + atom.

Polarizability: bigger, more e^- → e^- cloud distorts more easily → stronger LDF (I_2 solid, F_2 gas at 25 °C).

Shape: more surface contact → stronger LDF. n-pentane (bp 36 °C) > neopentane (bp 9.5 °C), same formula.

Dipole–dipole: δ+ end of 1 polar molecule attracts δ− end of the next. Needs a net molecular dipole.

H-bond: strong dipole–dipole; H covalently bonded to N, O or F attracted to a lone pair on another N, O or F.

Ion–dipole: ion attracts the oppositely charged end of a polar molecule → why salts dissolve in water.

IMF ≠ Bond

IMFs = attractions **between** particles; covalent bonds = **within** molecules. Boiling H_2O breaks IMFs only.

Coulomb's law behind all IMFs: bigger partial charges + closer particles → stronger attraction.

H-Bond Traps

CH_3OCH_3 (ether): no H on O → no H-bonds between its own molecules, but it can accept H-bonds from water.

CH_3F has F, but its H's are bonded to C, not F → no H-bonding; only dipole–dipole + LDF between molecules.

⚠ A C–H bond never H-bonds. The H must be bonded directly to N, O or F to count (CH_4, CH_3F: no H-bonds).

✓ FRQ: name the IMF on BOTH substances, say which is stronger + why, then link to the property.

3.1 Comparing IMFs & Properties

Ranking Strength

Typical order: ion–dipole > H-bond > dipole–dipole > LDF, but only when molecules are similar in size.

LDF can win: big nonpolar molecules can beat small polar ones. Octane (bp 126 °C) > water (bp 100 °C).

Same IMF type? Compare e^- count / size: He < Ne < Ar < Kr in bp; F_2 (g) < Cl_2 (g) < Br_2 (l) < I_2 (s).

Mixed IMFs: count all of them. CH_3OH has LDF + dipole–dipole + H-bonds; CH_4 has LDF only (bp −162 °C).

More H-bond sites → higher bp: H_2O (2 H + 2 lone pairs per molecule) > HF > NH_3 in bp (100, 20, −33 °C).

Pair	Higher bp	Why
H_2O vs H_2S	H_2O	H-bonds
HF vs HCl	HF	H-bonds
CH_3OH vs CH_4	CH_3OH	H-bond, polar
Cl_2 vs F_2	Cl_2	more e^- → LDF

Properties from IMFs

Stronger IMFs → ↑ bp, ↑ mp, ↑ ΔH_{vap}, ↑ viscosity, ↑ surface tension, ↓ vapor pressure (VP), ↓ volatility.

Vapor pressure: pressure of vapor above a liquid in a closed container; ↑ w/ T; weaker IMFs → higher VP.

Normal bp = T where VP = 1 atm. A low-VP liquid needs a higher T to reach 1 atm → it has a higher bp.

Volatile = weak IMFs → evaporates fast. Acetone evaporates faster than water b/c it has no H-bonds.

Evaporation is endothermic: fastest particles leave 1st → avg KE of the rest ↓ → the remaining liquid cools.

Surface tension: inward pull on surface particles; water's H-bonds → small insects can stand on water.

⚠ Bp compares IMFs, not bond strength. "H_2O has strong O–H bonds" is wrong reasoning for its high bp.

✓ Ranking MCQ: same size? → use polarity/H-bonds. Very different sizes? → LDF (e^- count) may dominate.

3.2–3.3 Solids, Liquids & Gases

Types of Solids

Solid	Particles	Props
Ionic	cations/anions	brittle, high mp
Covalent net.	atoms, bonds	very high mp
Molecular	molecules, IMFs	low mp, soft
Metallic	cations, e^- sea	conduct, malleable

Conductivity: ionic only when molten or (aq); metals always; molecular + most network solids never.

Covalent network: diamond, SiO_2, SiC. Melting means bonds are overcome completely → very high mp.

Graphite: covalent sheets w/ only LDF between layers → soft, slippery; delocalized e^- → it conducts.

Molecular solids (I_2, ice, sugar): melt by overcoming IMFs only → lower mp than ionic or network.

Ionic mp ↑ w/ charge + ↓ ion size (Coulomb): MgO (2852 °C) ≫ NaCl (801 °C) b/c +2/−2 vs +1/−1 charges.

Crystalline: ordered, repeating lattice. **Amorphous** (glass): no long-range order, softens over a range.

Phases at Particle Level

Solid: fixed positions, vibrate only. **Liquid**: close but slide past each other. **Gas**: far apart + fast.

Compressibility: gases yes (mostly empty space); liquids + solids barely (particles already touching).

Density: solid ≈ liquid ≫ gas. Ice is the exception: H-bonded open lattice → ice floats on liquid water.

Phase changes overcome IMFs; covalent bonds inside each molecule stay intact (H_2O stays H_2O as steam).

Gas vs liquid at same T: particles have similar KE; IMF strength decides whether they stay together.

⚠ Particle diagrams: liquid particles still touch, randomly arranged. Draw gas particles far apart, same size.

✓ "Explain mp difference": ID each solid type, name what must be overcome (ions, bonds, IMFs), compare.

3.4 Ideal Gas Law

Core Setups (Ref Sheet: PV = nRT)

Pick R by units: 0.08206 L·atm/(mol·K) for atm + L; 8.314 J/(mol·K) for energy. Always use T in **K**.

Combined law: $P_1V_1/T_1 = P_2V_2/T_2$ (n constant). Cancel whatever stays fixed (Boyle, Charles, Gay-Lussac).

Hold const.	Change	Result
n, T	P ↑	V ↓ (Boyle)
n, P	T ↑	V ↑ (Charles)
n, V	T ↑	P ↑
P, T	n ↑	V ↑ (Avogadro)

Molar mass: M = dRT/P (d in g/L). **Density**: d = PM/RT → heavier gas, higher P or lower T → denser.

STP: 273.15 K, 1 atm → 22.4 L/mol for an ideal gas. Use it only at STP; otherwise use PV = nRT.

Mixtures & Partial Pressure

Dalton: $P_{total} = P_A + P_B + \ldots$; each gas acts as if alone. At same V + T, P of each gas ∝ its moles.

Mole fraction: $P_A = X_A \cdot P_{total}$, $X_A = n_A/n_{total}$. 1 mol He + 3 mol Ar at 8 atm → $P_{He} = 0.25 \times 8 = 2$ atm.

Gas over water: $P_{gas} = P_{total} - P_{H_2O}$ (VP of water at that T is given), then plug P_{gas} into PV = nRT to get n.

Gas Stoichiometry

Path: P, V, T → n (PV = nRT) → mole ratio from the balanced eq → mass or V of the other species asked for.

Same T + P: volume ratio = mole ratio. $N_2 + 3H_2 \rightarrow 2NH_3$ means 1 L N_2 needs 3 L H_2 → makes 2 L NH_3.

⚠ °C in PV = nRT is the #1 error. 25 °C = 298 K. Also convert mL → L + torr/mmHg → atm (760 = 1 atm).

✓ Show units on every line of a gas FRQ; rubrics give credit for setup even if the arithmetic slips.

3.5–3.6 Kinetic Molecular Theory & Real Gases

KMT Assumptions (Ideal)

Particles have negligible volume vs container; move randomly in straight lines between collisions.

No attractions or repulsions between particles; collisions are elastic (total KE is conserved in each).

Avg KE ∝ T (K): all gases at the same T have the same avg KE, regardless of their mass. KE = ½mv².

Same KE, diff mass → lighter particles move faster on avg: H_2 fastest, then He, … at the same T (½mv² fixed).

Pressure = particle collisions w/ walls. ↑ T → faster + more forceful hits → ↑ P (at constant V + n).

Maxwell–Boltzmann Plots

↑ T → curve flattens + spreads right; peak lower; more particles at high speed. Area stays the same.

Read the plot: fraction w/ speed above a cutoff = area to its right; ↑ T → that area (fast fraction) grows.

Lighter gas (same T) → peak farther right + broader. Heavier gas → narrow, tall peak at low speed.

Real Gases

Deviate most at **high P** + **low T**: particles close together + slow → attractions + particle volume matter.

Factor	Effect	vs ideal
IMF attractions	softer wall hits	$P_{real} < P_{ideal}$
Particle volume	less free space	$V_{real} > V_{ideal}$

Most ideal: small, nonpolar, weak IMFs (He, H_2). **Least**: polar, large, H-bonding (NH_3, H_2O vapor, HF).

Why low T: slow particles can't escape attractions → gas acts less ideal + may condense to a liquid.

⚠ Same T → same avg KE, NOT same avg speed. Mixing these up loses MCQ + FRQ points every year.

✓ "Why does the real gas deviate?": name the IMF + say particles attract → fewer/softer wall collisions.

3.7–3.8 Solutions & Representations

Solutions & Mixtures

Pure substance: fixed composition. **Mixture**: variable proportions; components keep their identity.

Solution = homogeneous mixture: uniform throughout, can't be filtered; solute dissolved in solvent.

Solute = what dissolves (lesser amount); **solvent** = what does the dissolving (water in all aq solutions).

Heterogeneous: visible regions, varying composition (sand + water, oil + water). Separable by filtering.

Concentration

Molarity M = mol solute / L **solution** (not L solvent). 0.50 mol NaCl in 250 mL of solution → 2.0 M NaCl.

Dilution: $M_1V_1 = M_2V_2$ b/c moles of solute stay the same. Add water → V ↑, M ↓. 10 mL 6 M → 60 mL of 1 M.

Ion concentration: 0.10 M $CaCl_2$ → [Ca^{2+}] = 0.10 M, [Cl^-] = 0.20 M. Multiply by each ion's subscript.

Make a solution: weigh solute, dissolve in some solvent in a volumetric flask, then fill to the line + mix.

Glassware	Use
Volumetric flask	exact final volume
Buret / pipet	exact delivered vol.
Beaker	rough volume only

Particle Diagrams

Ions separate in water: draw Na^+ + Cl^- apart, never NaCl pairs. Keep the ratio ($CaCl_2$ → 2 Cl^- per Ca^{2+}).

Water orientation: O end (δ−) faces cations; H ends (δ+) face anions. Molecular solutes stay intact.

Molecular solute (sugar, C_2H_5OH): draw whole molecules, H-bonded to water, never split into ions.

⚠ Drawing water's H toward a cation (or O toward an anion) is a common lost point on particle FRQs.

✓ Diagram FRQ: show correct ion ratio, ion size, water orientation + a key. Count particles twice.

3.9–3.10 Separation & Solubility

Chromatography

Idea: components split by relative attraction to the **mobile phase** (solvent) vs the **stationary phase** (paper).

R_f = distance traveled by spot ÷ distance traveled by solvent front. Always 0 to 1, no units; measure from origin.

More attracted to stationary phase → moves less → smaller R_f. More attracted to solvent → larger R_f.

Polar paper + nonpolar solvent → polar dyes stick, low R_f.

Chromatogram → infer relative polarity.

Method	Separates by	Example
Paper/TLC	polarity, IMFs	ink dyes
Column	polarity, IMFs	plant pigments
Distillation	bp (IMFs)	ethanol + water
Filtration	phase/size	sand + water

Distillation: weaker IMFs → lower bp → vaporizes + collects 1st. Ethanol (bp 78 °C) comes off before water.

Solubility

Like dissolves like: solute–solvent IMFs must be comparable to the solute–solute + solvent–solvent ones.

Ionic + water: ion–dipole replaces ionic attractions → many salts dissolve. NaCl won't dissolve in hexane.

Polar + polar: CH_3OH mixes w/ water (H-bonds). **Nonpolar + nonpolar**: I_2 dissolves in CCl_4 via LDF.

Long nonpolar tail ↓ water solubility: methanol mixes fully; octanol barely dissolves (big LDF tail).

Solubility ↓ when new solute–solvent IMFs are weaker than those broken (oil + water stay separate).

⚠ "Polar dissolves polar" alone won't earn the point. Name the IMFs formed between solute + solvent.

✓ Separation FRQ: link each result (R_f value, collection order) to a specific IMF difference between parts.

3.11–3.13 Spectroscopy, Photons & Beer's Law

EM Spectrum & Molecules

Region	Energy	Causes
Microwave	low	rotation
Infrared	medium	bond vibration
UV/visible	high	e^- transitions

Energy order: radio < microwave < IR < visible < UV < X-ray. Shorter λ → higher ν → higher E per photon.

Absorb a photon → molecule's energy ↑ by exactly E_{photon}; **emit** → its energy ↓ by that same amount.

Photons (Ref Sheet: E = hν, c = λν)

Chain: λ → ν = c/λ → E = hν. Convert nm → m 1st (1 nm = 10^{-9} m). For per mole of photons: × 6.022×10^{23}.

E.g.: λ = 500 nm → ν = 6.0×10^{14} s^{-1} → E ≈ 4.0×10^{-19} J per photon (≈ 240 kJ/mol), visible green light.

E ∝ ν, E ∝ 1/λ: halve λ → double E. Brighter light = more photons, not more energy per photon.

Beer–Lambert Law (A = εbc)

A absorbance (no units), **ε** molar absorptivity ($M^{-1}cm^{-1}$), **b** path length (cm, often 1), **c** molarity (mol/L).

Calibration curve: A vs c for standards → straight line through origin; slope = εb → read unknown c.

Set λ to λ_{max} (max absorbance) → most sensitive. A solution absorbs the complement of the color you see.

Errors: fingerprints/bubbles on cuvette → A too high → c too high. Water left in cuvette → c too low.

Beer's law limits: line holds best at low c; very high A → dilute the sample, measure again, then × dilution factor.

⚠ ε depends on the substance + λ, not on concentration. Changing c changes A, never ε or b.

✓ Error FRQ: say if A goes ↑ or ↓, then carry it through A = εbc to say if calculated c is high or low.