The Simple AP Chemistry Unit 3 Cheat Sheet

3.1 Intermolecular Forces (IMFs)

IMF Types

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

London dispersion (LDF): brief dipoles attract each other. Found in **every** molecule + atom.
Polarizability: more e^- → bigger, squishier e^- cloud → stronger LDF (I_2 is solid, F_2 is a gas).
Dipole–dipole: δ+ end of 1 polar molecule attracts the δ− end of the next molecule.
H-bond: H bonded to N, O or F is attracted to a lone pair on another N, O or F atom.
Ion–dipole: an ion attracts the opposite end of a polar molecule → why salts dissolve in water.

IMF ≠ Bond

IMFs act **between** molecules; covalent bonds hold atoms together **within** 1 molecule.
Boiling water overcomes IMFs only. H_2O molecules stay whole as steam.

H-Bond Trap

CH_3F has F, but its H's are on C, not F → no H-bonds; just dipole–dipole + LDF.

⚠ A C–H bond never H-bonds. The H must be bonded directly to N, O or F atoms.

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

3.1 Comparing IMFs & Properties

Ranking Strength

Usual order: ion–dipole > H-bond > dipole–dipole > LDF, for molecules of similar size.
LDF can win: big nonpolar molecules can beat small polar ones. Octane (bp 126 °C) > water.
Same IMF type? More e^- → higher bp: He < Ne < Ar; F_2 (gas) < Cl_2 (gas) < Br_2 (liquid) < I_2 (solid).
Mixed IMFs: list all of them. CH_3OH has LDF, dipole–dipole + H-bonds; CH_4 has LDF only.

Pair	Higher bp	Why
H_2O vs H_2S	H_2O	H-bonds
HF vs HCl	HF	H-bonds
Cl_2 vs F_2	Cl_2	more e^-, LDF

Properties from IMFs

Stronger IMFs → higher bp + mp, more viscous, more surface tension, lower vapor pressure.
Vapor pressure (VP): push of vapor above a liquid. ↑ w/ T; weaker IMFs → higher VP.
Normal bp: the T where VP = 1 atm. Low VP liquid → needs a higher T → higher bp.
Volatile = weak IMFs → evaporates fast (acetone vs water: no H-bonds in acetone).
Evaporation cools: fastest particles leave 1st, so the avg KE of the liquid left drops.
Surface tension: water's H-bonds pull its surface tight, so small insects can stand on it.

⚠ A high bp comes from strong IMFs, not strong bonds. "Strong O–H bonds" is the wrong reason.

✓ Ranking tip: same size → compare polarity + H-bonds. Very different size → check e^- count.

3.2–3.3 Solids, Liquids & Gases

Types of Solids

Solid	Held by	Props
Ionic	ions	brittle, high mp
Covalent net.	covalent bonds	very high mp
Molecular	IMFs	low mp, soft
Metallic	e^- sea	conduct, bend

Conducts? Ionic only when melted or dissolved; metals always; molecular solids never.
Covalent network (diamond, SiO_2): must break covalent bonds to melt → very high mp.
Molecular solids (I_2, ice, sugar): melt by overcoming IMFs only → low mp.
Crystalline: ordered, repeating pattern.
Amorphous (glass): no long-range order.
Ionic mp: bigger charges + smaller ions → higher mp (MgO 2852 °C ≫ NaCl 801 °C).

Phases at Particle Level

Solid: particles fixed, vibrate. **Liquid**: touching but sliding past. **Gas**: far apart + fast.
Compressible? Gases yes (mostly empty space); liquids + solids barely (particles touch).
Ice floats: H-bonds hold water in an open lattice, so solid water is less dense than liquid.
Phase changes overcome IMFs only; the molecules themselves don't break apart.

⚠ In particle diagrams, liquid particles still touch. Only gas particles are drawn far apart.

✓ Explain mp: name each solid type + what must be overcome (ions, bonds or IMFs).

3.4 Ideal Gas Law

Core Setups (Ref Sheet: PV = nRT)

R = 0.08206 L·atm/(mol·K) for atm + L. T in **kelvin**.
Combined law: $P_1V_1/T_1 = P_2V_2/T_2$ (n constant). Cancel any variable that stays the same.

Hold const.	Change	Result
n, T	P ↑	V ↓
n, P	T ↑	V ↑
n, V	T ↑	P ↑
P, T	n ↑	V ↑

Molar mass: M = dRT/P (d = density in g/L).
Density: d = PM/RT → a heavier gas, higher P or lower T each give a denser gas sample.
STP (273.15 K, 1 atm): 1 mol gas = 22.4 L.

Mixtures & Partial Pressure

Dalton's law: $P_{total} = P_A + P_B + \ldots$
Mole fraction: $P_A = X_A \cdot P_{total}$. 1 mol He + 3 mol Ar at 8 atm → $P_{He} = 0.25 \times 8 = 2$ atm.
Over water: $P_{gas} = P_{total} - P_{H_2O}$.

Gas Stoichiometry

Path: PV = nRT → moles → mole ratio from the balanced equation → grams or liters asked for.
Same T + P: volume ratio = mole ratio ($N_2 + 3H_2 \rightarrow 2NH_3$: 1 L N_2 reacts w/ 3 L H_2 → 2 L NH_3).

⚠ Using °C in PV = nRT is the #1 error. 25 °C = 298 K. Also mL → L + torr → atm.

✓ Show units on every step of a gas FRQ; the setup earns points even if math slips.

3.5–3.6 KMT & Real Gases

KMT Assumptions (Ideal)

Particles are tiny compared w/ the container + move randomly in straight lines.
No attractions between particles; collisions are elastic (no KE is lost overall in a collision).
Avg KE ∝ T (K): all gases at the same T have the same avg KE, whatever their mass.
Lighter = faster: at the same T, H_2 moves faster on avg than O_2 (same KE, less mass).
Pressure comes from wall collisions. ↑ T → faster, harder hits → ↑ P (if V stays the same).

Maxwell–Boltzmann Plots

↑ T → curve spreads right + flattens; the peak gets lower. Total area stays the same.
Lighter gas (same T) → peak farther right + broader than the curve for a heavier gas.

Real Gases

Deviate most at **high P** + **low T**: particles are close + slow, so IMFs + volume matter.

Factor	Effect
IMF attractions	$P_{real} < P_{ideal}$
Particle volume	$V_{real} > V_{ideal}$

Most ideal: small, nonpolar gases (He, H_2). **Least**: polar, H-bonding gases (NH_3, H_2O).

⚠ Same T → same avg KE, NOT the same avg speed. Lighter particles move faster.

✓ Explain deviation: name the IMF + say attractions → fewer, softer wall hits.

3.7–3.8 Solutions & Representations

Solutions & Mixtures

Pure substance: fixed makeup. **Mixture**: parts in any ratio + each keeps its identity.
Solution = homogeneous mixture: same all the way through; can't be filtered out.
Solute = what dissolves; **solvent** = what does the dissolving (water in any aq solution).

Concentration

Molarity M = mol solute / L **solution**. 0.50 mol NaCl in 250 mL of soln → 2.0 M.
Dilution: $M_1V_1 = M_2V_2$ (moles of solute stay the same). Adding water → M goes down.
Ions: 0.10 M $CaCl_2$ → 0.20 M Cl^-.

Glassware	Use
Volumetric flask	exact final volume
Buret / pipet	exact volume added
Beaker	rough volume only

Particle Diagrams

Ions separate in water: draw Na^+ + Cl^- apart, never as NaCl pairs. Keep the ratio.
Water orientation: the O end (δ−) faces cations, while the H ends (δ+) face anions.
Molecular solutes (sugar): draw whole molecules; they stay intact + don't split into ions.

⚠ Pointing water's H toward a cation is a common lost point on particle-diagram FRQs.

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

3.9–3.10 Separation & Solubility

Chromatography

Idea: parts split by attraction to the **mobile phase** (solvent) vs **stationary phase** (paper).
R_f = spot distance ÷ solvent front distance, both measured from the start line. Always 0–1.
More attracted to the paper → moves less → smaller R_f. Use this to compare polarity.

Method	Separates by
Paper / column	polarity (IMFs)
Distillation	boiling point
Filtration	solid vs liquid

Distillation: weaker IMFs → lower bp → boils off 1st (ethanol, bp 78 °C, before water).

Solubility

Like dissolves like: solute + solvent need similar IMFs for the solute to dissolve well.
Salts in water: ion–dipole attractions pull ions out of the lattice. NaCl won't dissolve in hexane.
Polar + polar: CH_3OH mixes w/ water (H-bonds).
Nonpolar + nonpolar: I_2 in CCl_4 (LDF).
Long nonpolar tail → less soluble in water: methanol mixes fully; octanol barely does.
Oil + water stay separate: water's H-bonds are stronger than any attraction to oil.

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

✓ Separation FRQ: tie each result (R_f, order collected) to a specific IMF difference.

3.11–3.13 Spectroscopy & Beer's Law

EM Spectrum & Molecules

Region	Causes
Microwave	molecule rotation
Infrared (IR)	bond vibration
UV / visible	e^- transitions

Energy order: radio < microwave < IR < visible < UV < X-ray. Shorter λ → more energy.
Absorb a photon → molecule gains exactly that photon's energy; **emit** → loses it.

Photons (Ref Sheet: E = hν, c = λν)

Steps: ν = c/λ, then E = hν (nm → m first).
E.g.: λ = 500 nm → $\nu = 6.0\times10^{14}$ s^{-1} → E = hν ≈ 4.0×10^{-19} J per photon of green light.
Brighter light = more photons, not more energy per photon. Only λ (or ν) sets E.

Beer–Lambert Law (A = εbc)

A = absorbance, **ε** = molar absorptivity, **b** = path length (cm), **c** = molarity (mol/L).
Calibration curve: plot A vs c for known samples → straight line → read the unknown c.
Use λ_{max} (most-absorbed λ) → most sensitive.
Errors: fingerprints on cuvette → A too high → c too high. Extra water → c too low.

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

✓ Error FRQ: say if A goes up or down, then use A = εbc to say if c comes out high or low.