The Ultimate AP® Chemistry Unit 1 Cheat Sheet

1.1 Moles & Molar Mass

Why Moles

Can't count particles in lab → measure mass, convert to moles to link mass ↔ # of particles you can't see.
Avogadro's #: $N_A = 6.022\times10^{23}$ mol^{-1} (particles per mole), the bridge from moles → atoms, molecules or ions.
amu ↔ g/mol: avg mass of 1 particle in amu = molar mass in g/mol (H_2O: 18.02 amu ↔ 18.02 g/mol).
Mass of 1 particle = molar mass ÷ N_A: 1 H_2O molecule = 18.02 g ÷ 6.022×10^{23} = 2.99×10^{-23} g, tiny.
Ionic compounds: count formula units, not molecules (1 mol NaCl = N_A formula units = 2 mol ions).

Conversion Map

From → To	Do this
grams → moles	÷ molar mass (n = m/M)
moles → grams	× molar mass
moles → particles	× 6.022×10^{23}
particles → moles	÷ 6.022×10^{23}
molecules → atoms	× subscript

Worked Examples

Molar mass: CO_2 = 12.01 + 2(16.00) = 44.01 g/mol.
Mass → particles: 22.0 g CO_2 ÷ 44.01 g/mol = 0.500 mol × N_A = 3.01×10^{23} CO_2 molecules (3 sig figs).
Atoms inside: 0.500 mol CO_2 → 1.00 mol O atoms = 6.02×10^{23} O atoms; 9.03×10^{23} atoms total.
Compare samples: equal masses → the lighter molar mass has MORE moles & particles (1 g H_2 ≫ 1 g O_2).

⚠ Molecules ≠ atoms. 1 mol H_2O has 3 mol atoms (2 mol H + 1 mol O). Read what the Q counts first.
✓ FRQ: show the setup w/ units & sig figs; a bare number earns no calculation point, even if right.

1.2 Mass Spectra of Elements

Reading a Mass Spectrum

x-axis: mass (amu, m/z) → each peak = 1 isotope. **y-axis**: relative abundance (% or relative # of atoms).
Isotopes: same # protons, different # neutrons → different mass, same chemistry (same e^- config).
# of peaks = # of naturally occurring isotopes; tallest peak = most common isotope of the element.
How it works: atoms ionized → accelerated → deflected by a magnet; lighter ions bend more → sorted by mass.

Average Atomic Mass

Weighted average: Σ(isotope mass × fractional abundance).

Isotope	Mass (amu)	Abundance
^{35}Cl	34.97	75.78%
^{37}Cl	36.97	24.22%

Example: (34.97)(0.7578) + (36.97)(0.2422) = 35.45 amu → matches the periodic table's value for Cl.
Estimate first: the avg mass sits closest to the most abundant isotope; use it to cut wrong choices fast.
Br: ^{79}Br & ^{81}Br ≈ 50/50 → avg ≈ 80 (actual 79.90).

Working Backward

Find abundance: let x = fraction of isotope 1, (1 − x) = isotope 2 → solve avg = $m_1x + m_2(1-x)$ for x.
Example: Cu = 63.55; ^{63}Cu 62.93, ^{65}Cu 64.93 → 63.55 = 64.93 − 2x → x = 0.69 → 69% Cu-63.
Identify element: compute the avg mass from the spectrum → match it to the periodic table's atomic mass.
Not tested: spectra w/ several elements, or peaks from molecules or multiply charged ions. 1 element only.

⚠ No atom has the average mass. Cl atoms are ~35 or ~37 amu, never 35.45; 35.45 is a weighted mean.
✓ Only single-element spectra are tested. Convert % to decimals before multiplying (75.78% → 0.7578).

1.3 Elemental Composition

Pure Substances

Molecular (e.g., H_2O) vs. **formula unit** (ionic, e.g., NaCl: fixed ion ratio in a lattice, no discrete molecules).
Law of definite proportions: a compound's mass ratio of elements is always the same (H_2O is always 11.2% H).

Mass % = (mass of element in 1 mol ÷ molar mass) × 100.
Example: H in H_2O = 2(1.008)/18.02 = 11.2%.

Empirical Formula Steps

Step	Action
1. Assume 100 g	% → grams
2. Grams → moles	÷ molar mass
3. Divide by smallest	Get mole ratio
4. Make whole #s	× 2, 3... if .5, .33

Example: 40.0% C, 6.7% H, 53.3% O → 3.33 : 6.65 : 3.33 mol → ÷ 3.33 → 1 : 2 : 1 → **CH_2O** (empirical).

Empirical formula: lowest whole-# ratio of atoms. **Molecular formula** = (empirical) × n, n a whole number.
n = molar mass ÷ empirical mass. CH_2O (30.03 g/mol) & M = 180.2 g/mol → n = 6 → $C_6H_{12}O_6$ (glucose).

Ratio Traps

1.5 → × 2; **1.33 or 1.67** → × 3; **1.25** → × 4. Don't round 1.5 to 2 or 1.33 to 1; multiply every ratio instead.
Same empirical formula, different compounds: CH_2O, $C_2H_4O_2$, $C_6H_{12}O_6$ all have equal mass %.
Combustion data: all C → CO_2, all H → H_2O; mol CO_2 = mol C, mol H_2O × 2 = mol H; O by mass difference.

⚠ Mass ratio ≠ mole ratio. Convert grams to moles BEFORE comparing or dividing by the smallest.
✓ Mass % alone gives only the empirical formula; you also need the molar mass to get the molecular one.

1.4 Composition of Mixtures

Pure Substance vs. Mixture

	Pure	Mixture
Particle types	1	2 or more
Composition	Fixed	Varies
Mass %	Constant	Sample-dependent
Example	NaCl, H_2O	Salt water, air

Elemental Analysis

Elemental analysis: measure an element's mass in a sample → find the relative # of atoms or the purity.
Compare to theory: measured mass % vs. the pure compound's % → tells if it's pure or what's mixed in.
Purity % = (mass of pure compound ÷ mass of sample) × 100. Pure = 100%; real samples fall below.

Worked Example

Sample: 10.0 g impure NaCl contains 3.50 g Cl.
Step 1: Cl in pure NaCl = 35.45/58.44 = 60.7%.
Step 2: mass NaCl = 3.50 g ÷ 0.607 = 5.77 g.
Step 3: purity = 5.77/10.0 = 57.7% (impurity has no Cl).

Reasoning w/ Impurities

Impurity w/ HIGHER % of element → measured % ↑; **lower %** (or none) → measured % ↓ vs. pure.
Mixture of 2 compounds: use the known total mass + element mass → 2 equations, 2 unknowns; solve.
Wet sample (not fully dried): extra water mass → calculated % of solute ↓; mass of sample too high.
Hydrate: heat to drive off water; mass lost = water → mol H_2O per mol salt ($CuSO_4\cdot5H_2O$: 5 mol water per mol).

⚠ A mixture can't have one formula. Only pure substances follow the law of definite proportions.
✓ Error Qs: say which direction the % shifts and WHY (extra mass or missing mass) in 1 sentence each.

1.5 Atomic Structure & Electron Config

Atom & Coulomb's Law

Nucleus: protons (+) & neutrons (0); **electrons** (−) outside.
Coulomb's law: $F \propto q_1q_2/r^2$. ↑ nuclear charge or ↓ distance → stronger hold on e^- → harder to remove.
Core e^- (inner shells) shield valence e^- → **effective nuclear charge** ≈ protons − core e^- (Na: 11 − 10 = +1).

Aufbau Order (Fill Low → High)

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p. Capacity: s 2, p 6, d 10.
Shell n holds $2n^2$ e^-: n = 2 → 2s + 2p = 8; n = 3 → 18.
Shorthand: [noble gas] + the rest, e.g. S = [Ne]$3s^23p^4$.
Blocks: s (groups 1–2), p (13–18), d (3–12) → read an atom's config straight from its table position.

Atom/Ion	Configuration
O (8 e^-)	$1s^22s^22p^4$
Na	[Ne]$3s^1$
Fe	[Ar]$4s^23d^6$
Fe^{3+}	[Ar]$3d^5$
Cl^-	[Ne]$3s^23p^6$ = [Ar]

Ions

Anions: add e^- to the next open subshell. **Cations**: remove e^- from the highest n first (outermost shell).
Transition metals: lose 4s before 3d (Fe → Fe^{2+} = [Ar]$3d^6$).
Isoelectronic: same e^- config (O^{2-}, F^-, Ne, Na^+, Mg^{2+}).
Not tested: quantum numbers; exceptions like Cr & Cu.

⚠ 4s fills before 3d but EMPTIES first when forming cations: Fe^{2+} = [Ar]$3d^6$, not [Ar]$4s^23d^4$.
✓ Explain w/ Coulomb's law: name the charge (protons) AND the distance (shell) for full FRQ credit.

1.6 Photoelectron Spectroscopy (PES)

Reading a PES Spectrum

x-axis: binding energy (MJ/mol) = energy to remove an e^- from that subshell. Often drawn HIGH → LOW, left to right.
Each peak = 1 subshell. **Peak height** ∝ # of e^- in it (ideally).
Farthest-left peak (highest energy) = 1s core e^-; the lowest-energy peak = the valence e^-, easiest to remove.
How it works: photon in → e^- out; binding energy = photon energy − the ejected e^-'s kinetic energy (KE).

Example: Sodium

Subshell	Rel. height	Energy
1s	2	Highest
2s	2	High
2p	6	Medium
3s	1	Lowest (0.50)

Read config from peaks: relative heights 2, 2, 6, 1 → $1s^22s^22p^63s^1$ = 11 e^- = Na. Sum heights = atomic # (neutral).

Comparing Spectra

Next element (more protons): same subshells shift to HIGHER binding energy → ↑ nuclear charge pulls harder.
New peak appears when a new subshell starts filling (e.g., Mg → Al adds a small 3p peak at low energy).
2s vs. 2p: the 2s peak sits at higher energy → 2s e^- penetrate closer to the nucleus than 2p.
Huge gap between peaks = new shell (n changes).
Valence peak energy ≈ 1st ionization energy (Na: 0.50 MJ/mol ≈ 496 kJ/mol). Lowest peak = IE_1.

⚠ Check the axis direction: binding energy usually INCREASES to the left on AP PES graphs (reversed x-axis).
✓ Shifted peaks → explain w/ Coulomb's law: more protons, same shell → stronger attraction → higher energy.

1.7 Periodic Trends

Trend Table

Property	Across →	Down ↓
Atomic radius	↓	↑
1st IE	↑	↓
Electronegativity	↑	↓
Electron affinity	↑ (more −)	↓ (usually)

Why (Coulomb's Law)

Across a period: +1 proton, same shell, little added shielding → ↑ Z_{eff} → e^- pulled closer, held tighter → radius ↓, IE ↑.
Down a group: new shell → valence e^- farther & more shielded → weaker attraction → radius ↑, IE ↓.
Electronegativity: pull on shared e^- in a bond; F highest (4.0). Noble gases: highest IE in each period.

Exceptions & Ions

Be → B: IE ↓ b/c B's 2p e^- is higher in energy than 2s. **N → O**: IE ↓ b/c O's paired 2p e^- repel each other.
Successive IE: big jump when removing a core e^-. Mg: 738, 1451, **7733** kJ/mol → 2 valence e^- → Mg^{2+}.
Cations smaller than the parent atom (lose a shell/↓ e^- repulsion); **anions** larger (↑ e^- repulsion).
Isoelectronic radii: same e^-, more protons → smaller. O^{2-} > F^- > Na^+ > Mg^{2+} (8, 9, 11, 12 protons pulling 10 e^-).
Predict missing data: use neighbors' values (e.g., estimate an element's radius between its group members).

⚠ "More shells" alone isn't enough on an FRQ: tie it to distance & attraction (Coulomb's law).
✓ Trend FRQ: compare protons (or Z_{eff}), shells/distance, then attraction → the property. 3 links.

1.8 Valence Electrons & Ionic Compounds

Valence Electrons Drive Bonding

Bonding likelihood: set by the attraction between valence e^- & the nuclei of the 2 elements bonding.
Metals (low IE) lose e^-; **nonmetals** (high EA, EN) gain e^- → oppositely charged ions → ionic compounds.

Typical Ion Charges

Group	Valence e^-	Ion
1	1	+1
2	2	+2
13 (Al)	3	+3
15	5	−3
16	6	−2
17	7	−1

Goal: noble-gas config. Na → Na^+ [Ne]; Cl → Cl^- [Ar].
Transition metals: several possible charges (Fe^{2+}, Fe^{3+}); they lose 4s e^- before 3d e^-. Charge shown by Roman numeral.

Same Column, Same Formula

Analogous compounds: same column → same formula. NaCl → KBr, LiF; MgO → CaS; Na_2O → K_2S, Li_2Se.
Predict formulas: balance charges. Al^{3+} + O^{2-} → Al_2O_3 (2×3 = 3×2 = 6); Ca^{2+} + N^{3-} → Ca_3N_2 (3×2 = 2×3).
Unknown element: forms XCl_2 → likely group 2 (X^{2+}) → its oxide is XO and its nitride is X_3N_2 (charges balance).
Metal + nonmetal → ionic; 2 nonmetals → covalent (Unit 2).
Same column, same chemistry: equal valence e^- → similar reactions (Na & K both + H_2O → MOH + H_2).

⚠ Ionic formulas are formula units (lowest ratio): MgO, not Mg_2O_2; $CaCl_2$, not Ca_2Cl_4. Always reduce the ratio.
✓ MCQ: unfamiliar element? Use its column to predict its ion charge & its compounds by analogy.