The Simple AP Chemistry Unit 4 Cheat Sheet

4.1, 4.3–4.4 Reactions & Diagrams

Physical vs Chemical Change

Chemical change: bonds break + form → new substances w/ new formulas (rusting, burning).
Physical change: same substance in a new form; only IMFs change (melting, boiling).

Evidence	Catch
Gas bubbles	boiling also bubbles
Precipitate	cooling can crystallize
Color change	dilution fades color
Temp change	dissolving too

Dissolving a salt: ionic bonds break, but no new substance forms → features of both.
Best proof of a chemical change: a new substance w/ new properties is identified.

Balanced Equations

Conservation: same # of each atom + same total charge on both sides of the equation.
Balance by changing coefficients, never subscripts (H_2O can't become H_2O_2).

Particle Diagrams

Count atoms: $2H_2 + O_2 \rightarrow 2H_2O$ has 4 H + 2 O atoms before AND after the reaction.
Leftovers: 6 H_2 + 2 O_2 → 4 H_2O + 2 H_2 left over. Draw the extra H_2 in the after box.

⚠ Atoms are conserved, molecules are not: 3 molecules can become 2 ($2H_2 + O_2 \rightarrow 2H_2O$).

✓ Diagram FRQ: count each element before + after, show leftover reactant + add a key.

4.2 Net Ionic Equations

Split vs Keep

Split into ions	Keep whole
Strong acids (aq)	Weak acids
Strong bases (aq)	Solids (s)
Soluble salts (aq)	Liquids, gases

Strong acids: HCl, HBr, HI, HNO_3, $HClO_4$, H_2SO_4.
Strong bases: group 1 hydroxides.
Always soluble: salts of Na^+, K^+, NH_4^+ + NO_3^-. That's all the AP expects you to know.

3-Step Method

1. Molecular: $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$, written w/ full formulas.
2. Complete ionic: split every strong electrolyte that's (aq) into its separate ions.
3. Net ionic: cross out **spectator ions** (Na^+, NO_3^-) → $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$.

Must-Know Patterns

Reaction	Net ionic
Strong acid + base	$H^+ + OH^- \rightarrow H_2O$
Weak acid + OH^-	$HA + OH^- \rightarrow A^- + H_2O$
Metal + ion	$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

Check: atoms + charges balance, + every species has its (s), (l), (g) or (aq) state label.

⚠ Never split a solid, a gas or a weak acid into ions. $CaCO_3(s)$ stays as $CaCO_3(s)$.

✓ If every ion cancels, no reaction happens ($NaCl(aq) + KNO_3(aq)$ → nothing forms).

4.5 Stoichiometry: The Mole Map

The Core Path

Grams A → mol A → mol B → grams B: ÷ molar mass, × mole ratio, × molar mass.
Mole ratio comes from the coefficients in the balanced equation, not subscripts.

Start w/	Get moles by
Mass (g)	n = m / M
Solution	n = M × V (in L)
Gas	n = PV / RT

Worked Examples

$2H_2 + O_2 \rightarrow 2H_2O$: 16.0 g O_2 = 0.500 mol O_2 → × 2 = 1.00 mol H_2O = 18.0 g of H_2O.
Solution: 50.0 mL of 0.200 M $AgNO_3$ = 0.0100 mol Ag^+ → makes 0.0100 mol AgCl.
Gas: find n from PV = nRT (T in K, V in L), then use the mole ratio the same way as before.

Good Habits

Units: write each conversion factor w/ units so they cancel on the way to the answer.
Sig figs: match the least precise measurement; round only at the very end of the problem.
Particles: moles × 6.022×10^{23} = number of molecules (0.250 mol → 1.51×10^{23}).
Mass ↔ formula: grams → moles of each element → ÷ smallest → empirical formula.

⚠ Mole ratios use moles, not grams. Convert grams to moles before using the ratio.

✓ Stoich FRQ: a clear setup w/ units earns credit even if the arithmetic slips later on.

4.5 Limiting Reactant & Yield

Find the Limiting Reactant

Shortcut: divide moles of each reactant by its coefficient. Smallest = limiting reactant.
E.g.: $2H_2 + O_2$ w/ 2.0 mol H_2 + 0.50 mol O_2: 2.0/2 = 1.0 vs 0.50/1 = 0.50 → **O_2 limits**.

Species	Start	Change	End
H_2	2.0	−1.0	1.0
O_2	0.50	−0.50	0
H_2O	0	+1.0	1.0

Excess left: start − used. Here 1.0 mol H_2 is left over after all the O_2 is used up.

Yield

Theoretical yield: the most product the limiting reactant can make (1.0 mol H_2O).
Percent yield = actual ÷ theoretical × 100. 15.3 g made ÷ 18.0 g possible = 85.0%.
Below 100%: product lost in transfer, side reactions or reaction not finished.
Above 100%: product still wet or impure, so the mass you weighed is too high to trust.

Lab Errors

Trace it: does the error make the measured mass go up or down? Follow it to the answer.
Spilled product → actual yield ↓ → percent yield ↓. Wet product → percent yield ↑.

⚠ The limiting reactant isn't always the one w/ less mass; the coefficients decide.

✓ Heat to constant mass: reweigh until 2 readings match → all water is gone.

4.6 Introduction to Titration

Setup & Terms

Titration: add a **titrant** of known M from a buret to an **analyte** of unknown M in a flask.
Equivalence point: the titrant added exactly reacts w/ all of the analyte (mole ratio).
End point: the indicator changes color; a good indicator does so near the equivalence point.

Glassware	Use
Buret	adds titrant (±0.01 mL)
Pipet	exact analyte volume
Flask	holds analyte

The Calculation

Steps: mol titrant = M × V(L) → × mole ratio → mol analyte → ÷ V(L) = analyte M. Show each.
E.g.: 20.00 mL of 0.100 M NaOH reacts w/ 25.00 mL HCl → 2.00×10^{-3} mol → 0.0800 M.
2 : 1 ratio: H_2SO_4 + 2NaOH, so mol H_2SO_4 = mol NaOH ÷ 2. 3.00×10^{-3} mol NaOH → 1.50×10^{-3} mol.

Technique & Errors

Rinse the buret w/ titrant: water left inside dilutes it → analyte M comes out too high.
Overshoot the end point → too much titrant used → analyte M comes out too high.
Air bubble in the buret tip → volume reads too high → analyte M comes out too high.

⚠ Equivalence ≠ pH 7 every time. Only strong acid + strong base is neutral there.

✓ Read the buret at eye level at the bottom of the meniscus, to 2 decimal places.

4.7–4.8 Reaction Types & Acid–Base

3 Main Reaction Types

Type	What moves
Acid–base	H^+ (a proton)
Redox	electrons (e^-)
Precipitation	ions form a solid

Precipitation: 2 soluble salts swap partners → a solid forms ($Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4$).
Combustion is redox: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$.
Redox clue: if any oxidation number changes, it's redox, whatever else it looks like.

Brønsted–Lowry Acids & Bases

Acid = gives H^+. **Base** = takes H^+. In water, H^+ exists as **H_3O^+** (hydronium ion).
Conjugate pair: 2 species that differ by 1 H^+, like NH_3/NH_4^+ or H_2O/OH^- in $NH_3 + H_2O$.
Amphoteric: can act as an acid or a base (H_2O, HCO_3^-). It depends on the partner.
Neutralization: acid + base → water + salt; the net ionic is $H^+ + OH^- \rightarrow H_2O$ for strong ones.

Strong vs Weak

Strong acids ionize fully (write H^+); **weak acids** barely do (keep HA whole, e.g., HF).
Polyprotic: H_2SO_4 + H_3PO_4 give up more than 1 H^+; count each H^+ in titration math.

⚠ Conjugate base of H_2O = OH^-; conjugate acid of H_2O = H_3O^+. Don't swap them.

✓ Label each conjugate pair; the 2 members always differ by exactly 1 H^+ (+ 1 unit of charge).

4.9 Oxidation Numbers & Redox

Oxidation Number Rules

Species	Ox. #
Element (Fe, O_2)	0
Monatomic ion	= its charge
O in compounds	−2 (usually)
H w/ nonmetals	+1

Sum rule: ox. #s add to 0 in a neutral compound or to the charge of a polyatomic ion.
E.g.: Mn in MnO_4^-: x + 4(−2) = −1, so x = **+7**. S in SO_4^{2-}: x + 4(−2) = −2, so S = **+6**.

Who Gets Oxidized?

OIL RIG: Oxidation Is Loss of e^- (ox. # ↑); Reduction Is Gain of e^- (ox. # ↓).
Oxidizing agent gets **reduced**; **reducing agent** gets **oxidized** (it gives away e^-).

$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$	Role
Zn: 0 → +2	oxidized
Cu^{2+}: +2 → 0	reduced

e^- lost = e^- gained in every redox reaction; that keeps the total charge balanced.

Redox Clues

Free element on 1 side + in a compound on the other → almost always a redox reaction.

⚠ Name the whole species as the agent: "MnO_4^- is the oxidizing agent," not "Mn."

✓ Justify: give the ox. # before + after for the element that changes, not just "gains e^-."

4.9 Half-Reactions & Exam Strategy

Half-Reactions

Oxidation half: $Fe^{2+} \rightarrow Fe^{3+} + e^-$ (electrons end up on the product side of the arrow).
Reduction half: $Cu^{2+} + 2e^- \rightarrow Cu$ (electrons are on the reactant side, b/c they're gained).
Balance in acid: balance main atoms, then O w/ H_2O, H w/ H^+, + charge w/ e^-, in that order.
Combine: multiply so the e^- match, add the 2 halves, then cancel the electrons.
E.g.: $MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$. Check: charge is +17 on each side.

Exam Moves

Asked	Do this
Net ionic	split strong (aq) only
Limiting reactant	mol ÷ coefficient
Titration M	mol ratio, then ÷ V
Redox ID	ox. # before + after

Units: change mL to L before M × V; get molar masses from the periodic table.
Claim + evidence: say what's true ("O_2 is limiting"), then show the numbers.
Redox titration: purple MnO_4^- is its own indicator; the end point = the 1st lasting pink color.
Sig figs + units on every final answer; match the data's precision (3 s.f. in → 3 s.f. out).

⚠ Electrons must cancel when you add half-reactions. Leftover e^- = mistake.

✓ The periodic table + equations sheet are given; you bring the setup + reasoning.