The Simple AP Chemistry Unit 2 Cheat Sheet

2.1 Types of Chemical Bonds

Bond Type by Electronegativity

Electronegativity (EN): how hard an atom pulls on shared e^-. ↑ across a period, ↓ down a group. F is highest (~4.0); Cs & Fr are lowest.

Bond	Between	Electrons
Ionic	Metal+nonmet.	e^- transferred
Polar cov.	Nonmetals	Shared unevenly
Nonpolar	Same EN	Shared evenly
Metallic	Metal+metal	e^- sea

ΔEN rule of thumb: ~0–0.4 nonpolar, ~0.4–1.8 polar, >~1.8 mostly ionic. Bonding is a continuum, not 3 separate boxes w/ hard edges.

Polar bond: less EN atom gets δ+, more EN atom gets δ− → **bond dipole** points toward δ− end.

Reading Bond Character

Ionic character ↑ as ΔEN ↑. CsF most ionic; even NaCl has slight covalent character.

Metal + nonmetal → ionic (NaCl, MgO). **Nonmetal + nonmetal** → covalent (CO_2, NH_3). **Metalloids** (Si, B) bond covalently (SiO_2, BF_3, $SiCl_4$).

⚠ Ionic compounds aren't molecules. NaCl is a lattice of ions: say "formula unit," not "molecule."

Exam Moves

Justify bond type w/ ΔEN or metal/nonmetal ID, never just "b/c it's ionic." Name both.

✓ FRQ "which bond is more polar?": cite both ΔEN values (H–F ≈ 1.8 vs H–Cl ≈ 1.0), then compare them directly in 1 line.

2.2 Bonds & Potential Energy

Potential Energy Curve

PE vs. internuclear distance: atoms far apart → PE = 0. Closer → nucleus–e^- attraction lowers PE. Too close → nucleus–nucleus repulsion → PE shoots up fast (the steep left wall of the curve).

Bond length = distance at the PE minimum (attraction & repulsion balanced).

Bond energy = depth of the well = energy needed to break the bond (in kJ/mol of bonds).

⚠ Breaking a bond ALWAYS absorbs energy (endothermic). Forming a bond ALWAYS releases energy (exothermic). No exceptions.

Bond Order, Length & Strength

Bond	Order	Length	Energy
C–C	1	Longest	Weakest
C=C	2	Middle	Middle
C≡C	3	Shortest	Strongest

↑ Bond order → shorter, stronger bond (more shared e^- pull the nuclei closer).

Bigger atoms → longer, weaker bonds: H–I > H–Br > H–Cl > H–F in bond length (down group 17).

Compare 2 curves: deeper well = stronger bond; minimum farther right = longer bond.

Coulomb's Law

Coulomb's law: $F \propto q_1q_2/r^2$. ↑ charge or ↓ distance → stronger attraction & deeper PE well.

✓ Curve FRQ: mark bond length at the minimum, bond energy as the well depth, and explain each region w/ attraction vs. repulsion.

2.3 Structure of Ionic Solids

Lattice Basics

Ionic solid: 3D lattice of alternating cations & anions; each ion ringed by opposite charges → max attraction, min repulsion, very stable.

Lattice energy (LE): energy to split 1 mol solid into gas ions. LE ∝ $(q_+ \cdot q_-)/(r_+ + r_-)$, Coulomb.

Compare Lattice Energy

Compare	Bigger LE	Why
NaCl vs MgO	MgO	+2/−2 charges
NaCl vs KCl	NaCl	Na^+ smaller
NaF vs NaCl	NaF	F^- smaller

Check charge first, then size. Charge product dominates: MgO (2×2) ≫ NaCl (1×1).

⚠ Compare ION sizes, not atom sizes. Na^+ is much smaller than Na; Cl^- is bigger than Cl.

Properties & Why

High mp/bp: strong Coulombic attraction through the whole lattice. ↑ LE → ↑ mp (MgO 2852 °C vs NaCl 801 °C, b/c of +2/−2 ions).

Brittle: shift a layer → like charges line up → repel → the crystal shatters, not bends.

Conductivity: solid NO (ions locked). Molten or (aq) YES (ions mobile). Ions carry charge, not e^-.

Melting/dissolving breaks ionic attractions between ions, not covalent bonds.

✓ Particle diagrams: alternate + and − ions, draw cations & anions to relative size, and show no gaps in the lattice at all (Cl^- > Na^+).

2.4 Structure of Metals & Alloys

Metallic Bonding

Electron sea model: lattice of metal cations in a sea of delocalized valence e^- shared by all.

Conductive: mobile e^- carry charge (solid & liquid). **Malleable & ductile**: cations slide, and the e^- sea re-forms around them, so no repulsion.

Bond strength ↑ w/ more valence e^- & smaller cation → mp of Mg (650 °C) > Na (98 °C).

Alloys

Alloy	Atoms	Example
Substitutional	Similar radii	Brass (Cu+Zn)
Interstitial	Small in gaps	Steel (Fe+C)

Substitutional: solute atom replaces a host atom (similar size, e.g. brass). Density sits between the 2 metals; the alloy stays malleable.

Interstitial: small atoms (C, N, B, H) fill holes in lattice → ↑ rigidity, strength & density; ↓ malleability & ductility (layers can't slide).

⚠ Interstitial alloys get harder & LESS malleable. "Stronger and more malleable" is a classic wrong answer choice on MCQs here.

Metal vs. Ionic Solid

Property	Metal	Ionic
Conducts solid	Yes	No
Conducts melted	Yes	Yes
Hit w/ hammer	Bends	Shatters

✓ Explain props at particle level: "delocalized e^- move freely" earns the point; "metals conduct" alone does not earn any credit.

2.5 Lewis Diagrams

5-Step Method

1. Count valence e^-: sum group #s; add 1 e^- per − charge, subtract 1 per + charge.

2. Skeleton: least EN atom central (never H). H & F always terminal. Single bonds to all.

3. Fill octets on the outer atoms, then put any leftover e^- on the central atom.

4. Central atom short? Turn outer lone pairs into double or triple bonds until it has 8.

5. Check: e^- drawn = step 1 count. Ions get [brackets] + the charge outside.

Typical bonding (FC 0): C → 4 bonds; N → 3 bonds + 1 lone pair; O → 2 bonds + 2 lone pairs; halogen → 1 bond + 3 lone pairs.

Worked Counts

Species	Valence e^-	Result
CO_2	4+2(6)=16	O=C=O
NH_4^+	5+4(1)−1=8	4 N–H, [ ]$^+$
NO_3^-	5+3(6)+1=24	1 N=O, 3 res.

Octet Exceptions

Incomplete octet: B (BF_3: 6 e^-), Be ($BeCl_2$: 4 e^-); stable w/o double bonds, so leave them.

Odd e^- (radicals): NO, NO_2 → one atom is stuck at 7 e^-; no full octet is possible for all.

Expanded octet: period 3+ central atom (P, S, Cl, Xe) → PCl_5 (10 e^-), SF_6 (12), XeF_4 (12).

⚠ Period 2 atoms (C, N, O, F) can NEVER exceed 8 e^-. No 2nd-row expanded octets.

2.6 Resonance & Formal Charge

Formal Charge (FC)

FC = valence e^- − (lone-pair e^- + ½ bonding e^-). Quick: FC = V − (dots + lines) on that atom.

Sum of FCs = overall charge (0 if neutral, −1 for NO_3^-, −2 for CO_3^{2-}). Use it to check.

Best structure: FCs closest to 0; any − FC on most EN atom; avoid like charges side by side.

Example: SCN^- best = $[S{=}C{=}N]^-$, w/ the −1 FC on N (N is more EN than S). Sum = −1.

FC is a tiebreaker, not a real charge.

CO_2 option	FC (O, C, O)	Verdict
O=C=O	0, 0, 0	✓ Best
O≡C–O	+1, 0, −1	× Charges

Resonance

Resonance: ≥2 valid Lewis structures that differ only in e^- placement (atoms fixed). The real species = the **hybrid** (an average).

Hybrid bond order = total bonds ÷ positions. NO_3^- & CO_3^{2-}: 4/3 ≈ 1.33. O_3: 3/2 = 1.5.

Evidence: all 3 N–O bonds in NO_3^- have equal length, between single & double.

Resonance ≠ isomers: isomers move atoms; resonance moves only e^- (lone pairs, π bonds).

⚠ Resonance structures do NOT flip back & forth. The species is always the one averaged hybrid; all equal bonds are 1 length.

✓ FRQ "which structure is best?": calculate FC for every atom in each, then cite the rule (FCs nearest 0, any − on the most EN atom).

2.7 VSEPR & Molecular Geometry

Count Electron Domains

Electron domain = 1 lone pair or 1 bond (single, double & triple each count as ONE).

Electron geometry uses all domains; **molecular shape** names the atoms only (H_2O: bent).

Domains	Lone pr	Shape	Angle
2	0	Linear	180°
3	0	Trig. plan.	120°
3	1	Bent	<120°
4	0	Tetrahed.	109.5°
4	1	Pyramidal	~107°
4	2	Bent	~104.5°
5	0	Trig. bipy.	90°,120°
5	1	Seesaw	<90°,<120°
5	2	T-shaped	<90°
5	3	Linear	180°
6	0	Octahedral	90°
6	1	Sq. pyram.	<90°
6	2	Sq. plan.	90°

⚠ Lone pairs repel more than bonding pairs → they squeeze bond angles: CH_4 109.5° > NH_3 107° > H_2O 104.5° (0, 1, 2 lone pairs).

Lone Pair Placement

5 domains: lone pairs go equatorial (seesaw SF_4, T-shaped ClF_3, linear XeF_2) to cut 90° repulsion.

✓ Draw the Lewis structure first. Never guess shape from formula: CO_2 is linear but SO_2 is bent.

2.7 Hybridization, σ/π Bonds & Polarity

Hybridization

Hybridization (orbital mixing): # domains on an atom = # orbitals mixed (4 → sp^3).

Domains	Hybrid	Examples
2	sp	CO_2, HCN, C_2H_2
3	sp^2	BF_3, C_2H_4, SO_2
4	sp^3	CH_4, NH_3, H_2O

Shortcut: triple or 2 doubles → sp; 1 double → sp^2; all single → sp^3 (count lone pairs too).

Sigma (σ) & Pi (π) Bonds

σ bond: head-on overlap, e^- on the bond axis. **π bond**: side-by-side p overlap, e^- above & below the axis. Every bond has exactly 1 σ.

Single = 1σ. **Double** = 1σ + 1π. **Triple** = 1σ + 2π.

Count: C_2H_4 → 5σ, 1π. HCN → 2σ, 2π. CO_2 → 2σ, 2π. N_2 → 1σ, 2π. Double = 1σ + 1π.

Molecular Polarity

Polar molecule = polar bonds + asymmetric shape (dipoles don't cancel): H_2O, NH_3, SO_2, CH_2Cl_2.

Nonpolar: symmetric shape w/ identical outer atoms → dipoles cancel: CO_2, BF_3, CH_4, SF_6, XeF_4.

Same shape, ≠ polarity: CCl_4 nonpolar vs $CHCl_3$ polar (both tetrahedral; unequal outer atoms).

⚠ Polar bonds ≠ polar molecule. CO_2 has polar C=O bonds but is nonpolar (linear, so the 2 bond dipoles cancel exactly, net = 0).

✓ Polarity FRQ: name the shape, then say whether the bond dipoles cancel and why.