# The Ultimate AP Chemistry Unit 2 Cheat Sheet



## 2.1 Types of Chemical Bonds

### Bond Type by Electronegativity

**Electronegativity (EN)**: atom's pull on shared e⁻. ↑ across a period, ↓ down a group. F highest (~4.0).
**Why EN trends**: ↑ $Z_{eff}$ & smaller radius → the nucleus pulls bonding e⁻ harder; noble gases rarely bond, so EN not listed.

| Bond | Between | e⁻ behavior |
|---|---|---|
| Ionic | Metal+nonmetal | Transferred→ions |
| Polar cov. | Nonmetals, ≠ EN | Shared unequally |
| Nonpolar cov. | Same/similar EN | Shared equally |
| Metallic | Metal+metal | Delocalized sea |

**ΔEN rule of thumb**: ~0–0.4 nonpolar, ~0.4–1.8 polar, >~1.8 mostly ionic. Bonding is a continuum, not 3 boxes.
**Polar bond**: less EN atom gets δ+, more EN atom gets δ− → **bond dipole** points toward the δ− end (H→Cl).

### Reading Bond Character

**Ionic character** ↑ as ΔEN ↑. CsF most ionic; even NaCl has slight covalent character. No bond is 100% ionic.
**Metal + nonmetal** → ionic (NaCl, MgO). **Nonmetal + nonmetal** → covalent ($CO_2$, $NH_3$). **Metalloids** (Si, B) bond covalently w/ nonmetals ($SiO_2$, $BF_3$), never ionically.
**Polyatomic ions**: covalent bonds inside (N–O in $NO_3^-$), ionic attraction between ions → $NH_4NO_3$ has both.
**Pure covalent**: identical atoms ($H_2$, $Cl_2$, $O_2$) → ΔEN = 0, e⁻ shared evenly, no bond dipole at all; C–H is nearly nonpolar.

⚠ 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 elements & their types.
**Most polar bond** in a set = biggest ΔEN (H–F > H–O > H–N > H–C); rank by the partner atom's EN value.

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

## 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 ↑.
**Bond length** = distance at the PE minimum (attraction & repulsion balanced); read it off the x-axis.
**Bond energy** = depth of the well = energy needed to break the bond; read it off the y-axis (kJ/mol).
**2 curves, 1 graph**: H–F well is deeper & its minimum sits closer to r = 0 than H–Cl's → stronger, shorter.
**Graph reading**: deeper well → stronger bond. Minimum farther right → bigger atoms, longer bond.

⚠ Breaking a bond ALWAYS absorbs energy (endothermic). Forming a bond ALWAYS releases energy.

### 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 2 nuclei closer together): N≡N very strong.
**Bigger atoms** → longer, weaker bonds: H–I > H–Br > H–Cl > H–F in length (bond energy runs opposite).
**Diatomics**: $H_2$ (74 pm) is shorter than $Cl_2$ (199 pm) b/c Cl has more e⁻ shells → bigger atomic radius.

### Coulomb's Law

**Coulomb's law**: $F \propto q_1q_2/r^2$. ↑ charge or ↓ distance → stronger attraction & deeper PE well on the graph.
**Charge beats size**: doubling both charges = 4× the force; small size shifts change it far less than charge.

✓ Curve FRQ: mark bond length at the minimum, bond energy as the well depth, and explain each region w/ attraction vs. repulsion (far right: no interaction).

## 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 → a rigid, low-energy array.
**Lattice energy (LE)**: energy to split 1 mol solid into gas ions. LE ∝ $(q_+ \cdot q_-)/(r_+ + r_-)$: bigger charges & smaller ions → bigger LE → higher melting point & a harder crystal.
**Repeating pattern**: in NaCl each $Na^+$ touches 6 $Cl^-$ & each $Cl^-$ touches 6 $Na^+$ (a cubic array; no NaCl molecules).
**Formula from charges**: the lattice is neutral → the ion ratio balances charge ($CaF_2$, $Al_2O_3$, $Mg_3N_2$).

### 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 |
| MgO vs CaO | MgO | $Mg^{2+}$ smaller |

**Check charge first**, then size. Charge product dominates: MgO (2×2 = 4) ≫ NaCl (1×1 = 1) in LE.
**Same cation, bigger anion**: LE & mp ↓ from KF → KCl → KBr → KI as the anion radius ↑ down group 17.

⚠ Compare ION sizes, not atom sizes. $Na^+$ is much smaller than Na; $Cl^-$ is bigger than Cl (extra e⁻).

### Properties & Why

**High mp/bp**: strong Coulombic attraction through whole lattice. ↑ LE → ↑ mp (MgO 2852 °C vs NaCl 801 °C).
**Brittle**: shift a layer → like charges line up → repel → the crystal cleaves along a plane; it doesn't bend.
**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 inside them (e.g. in $NO_3^-$).

✓ Particle diagrams: alternate + and − ions, draw cations & anions to relative size, show no gaps in the lattice.

## 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 of the cations.
**Conductive**: mobile e⁻ carry charge (solid & liquid).
**Malleable & ductile**: cations slide past each other and the e⁻ sea re-forms around them, so no repulsion.
**Bond strength** ↑ w/ more valence e⁻ & a smaller cation → mp of Mg (650 °C) > Na (98 °C); 2 e⁻ vs 1 e⁻ per atom.
**Liquid metals** still conduct: the cations & e⁻ sea stay intact on melting (Hg is a liquid conductor at 25 °C).

### Alloys

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

**Substitutional**: solute atom replaces host atom (similar size). Density between the 2 metals; stays malleable.
**Interstitial**: small atoms (C, N, B, H) fill holes in the lattice → ↑ rigidity, strength & density; ↓ malleability & ductility (layers can't slide past each other easily, so it's harder).
**Alloy = mixture**, not a compound: composition varies; its props differ from the pure metals it's made of.

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

### Metal vs. Ionic Solid

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

**Particle diagram**: equal-sized circles mixed for substitutional; small dots in the gaps for interstitial (steel).

✓ Explain props at particle level: "delocalized e⁻ move freely" earns the point; "metals conduct" does not.

## 2.5 Lewis Diagrams

### 5-Step Method

**1. Count valence e⁻**: sum group #s; add 1 e⁻ per − charge, subtract 1 e⁻ per + charge ($NH_4^+$: 5 + 4 − 1 = 8).
**2. Skeleton**: least EN atom central (never H). H & F always terminal. Single bonds from center to all outer atoms.
**3. Fill octets** on the outer atoms, then put any leftover e⁻ on the central atom as lone pairs (in pairs of 2).
**4. Central atom short?** Turn outer lone pairs into double/triple bonds until it has 8 e⁻ ($CO_2$: 2 double bonds).
**5. Check**: e⁻ drawn = step 1 count. Ions get [brackets] + the charge written outside, top right: $[NO_3]^-$.
**Central atom**: $H_2O$ → O, $NH_3$ → N, HCN → C, $ClO_3^-$ → Cl, $SO_2$ → S. Usually the atom written once, least EN.
**Oxyacids**: H bonds to O, never to the central atom ($HNO_3$ → H–O–N; $H_2SO_4$ → 2 S–O–H units; the acidic H's).
**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. Use it to check your drawings fast.

### 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. |
| $SO_4^{2-}$ | 6+4(6)+2=32 | S central |

### Octet Exceptions

**Incomplete octet**: B ($BF_3$: 6 e⁻), Be ($BeCl_2$: 4 e⁻); stable w/o double bonds, so don't add them to fix the octet.
**Odd e⁻ (radicals)**: NO, $NO_2$ → one atom stuck at 7 e⁻.
**Expanded octet**: period 3+ central atom (P, S, Cl, Xe) → $PCl_5$ (10 e⁻), $SF_6$ (12 e⁻), $XeF_4$ (12 e⁻); empty d space allows it.
**Lone pairs count**: put the extra e⁻ on the central atom of $XeF_4$, $ICl_4^-$ and $I_3^-$, after the outer octets are full.

⚠ Period 2 atoms (C, N, O, F) can NEVER exceed 8 e⁻. No 2nd-row expanded octets, ever; recount if you get 10.

## 2.6 Resonance & Formal Charge

### Formal Charge (FC)

**FC** = valence e⁻ − (lone-pair e⁻ + ½ bonding e⁻). Quick: FC = V − (dots + lines) around that atom only.
**Sum of FCs** = overall charge (0 if neutral, −1 for $NO_3^-$).
**FC vs. oxidation #**: FC splits bonding e⁻ evenly; ox. # gives them all to the more EN atom in the bond.
**Best structure**: FCs closest to 0; any − FC on the 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). FCs: S 0, C 0, N −1; sum = −1.

| $CO_2$ 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). Real species = **hybrid** (average).
**Hybrid bond order** = total bonds ÷ positions. $NO_3^-$ & $CO_3^{2-}$: 4/3 ≈ 1.33. $O_3$: 3/2 = 1.5 per O–O bond, all equal.
**Evidence**: all 3 N–O bonds in $NO_3^-$ have equal length, between a single & a double bond's length.
**Resonance ≠ isomers**: isomers move atoms; resonance moves only e⁻ (lone pairs, π bonds), never atoms.
**Draw it**: show every structure w/ ↔ between them; move only lone pairs & π bonds, never atoms.

⚠ Resonance structures do NOT flip back & forth. The species is always the one averaged hybrid.

### Limits of the Lewis Model

**Lewis model** can't show delocalized e⁻ in 1 picture (needs resonance) or explain odd-e⁻ species.
**FC is a tiebreaker**, not real charge; use it only to choose among valid Lewis diagrams for the same atoms.

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

## 2.7 VSEPR & Molecular Geometry

### Count Electron Domains

**Electron domain** = 1 lone pair or 1 bond (single, double & triple each count as ONE domain around the atom).
**Electron geometry** uses all domains; **molecular shape** names the atoms only (lone pairs invisible).

| Domains | Lone pr | Shape | Angle |
|---|---|---|---|
| 2 | 0 | Linear | 180° |
| 3 | 0 | Trig. planar | 120° |
| 3 | 1 | Bent | <120° |
| 4 | 0 | Tetrahedral | 109.5° |
| 4 | 1 | Trig. pyram. | ~107° |
| 4 | 2 | Bent | ~104.5° |
| 5 | 0 | Trig. bipyr. | 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. planar | 90° |

⚠ Lone pairs repel more than bonding pairs → squeeze angles: $CH_4$ 109.5° > $NH_3$ 107° > $H_2O$ 104.5°.

### Lone Pair Placement

**5 domains**: lone pairs go equatorial (seesaw $SF_4$, T-shaped $ClF_3$, linear $XeF_2$) → less 90° repulsion.
**6 domains**: 2 lone pairs sit opposite each other (180° apart) → square planar $XeF_4$, $ICl_4^-$ (90° bond angles).
**Exact angle not asked?** Write "less than 109.5°" for $NH_3$ & $H_2O$; that earns the FRQ point, w/ the lone-pair reason.

✓ 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**: # domains on an atom = # orbitals mixed (1 s + p's): 4 domains → 1 s + 3 p = 4 $sp^3$ orbitals.

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

**Each atom** has its own hybridization: in $CH_3CHO$, the $CH_3$ carbon is $sp^3$, the C=O carbon is $sp^2$.
**Shortcut**: triple or 2 doubles → sp; 1 double → $sp^2$; all single → $sp^3$ (count lone pairs as domains too).
**d-orbital hybrids** (5 & 6 domains) aren't assessed; just know those shapes & angles from the VSEPR table.

### 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.
**Single** = 1σ. **Double** = 1σ + 1π. **Triple** = 1σ + 2π.
**Count**: $C_2H_4$ → 5σ, 1π. HCN → 2σ, 2π. $CO_2$ → 2σ, 2π.
**π e⁻** can be delocalized over several atoms → resonance ($O_3$, $NO_3^-$, benzene's ring); σ e⁻ stay between 2 atoms.

⚠ Every bond has exactly ONE σ bond. The extra bonds in double/triple bonds are π bonds (C≡N: 1σ + 2π).

### 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 → net dipole).

⚠ Polar bonds ≠ polar molecule. $CO_2$ has polar C=O bonds but is nonpolar (linear, dipoles cancel).

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