AP Chemistry — Unit 2 Review

Compound Structure and Properties: bonding, energy, alloys, Lewis structures, resonance, and molecular geometry

Source video: Jeremy Krug, AP Chemistry Unit 2 review (UltimateReviewPacket.com)

1. Ionic vs. Covalent Bonds

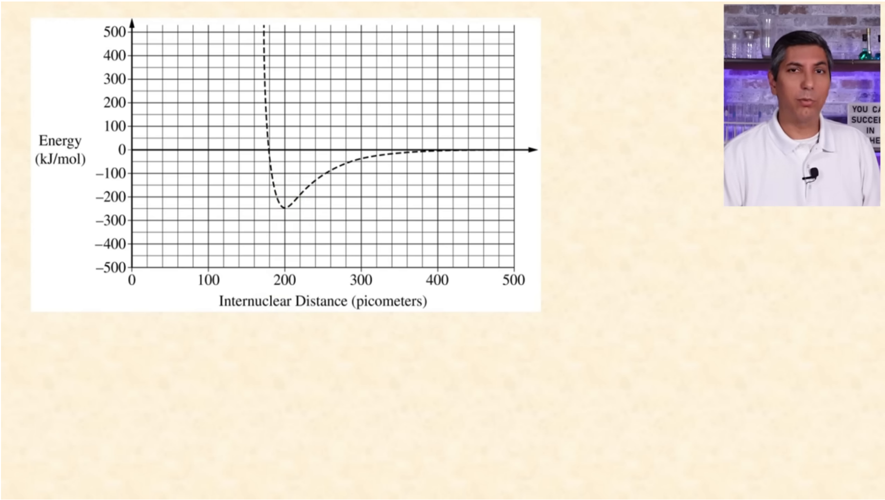
Ionic bonds form between a metal and a non-metal; covalent bonds form between two non-metals. Each bond type produces compounds with distinct physical properties:

Property	Ionic Compounds	Covalent Compounds
Bond formed between	Metal + non-metal	Two non-metals
Melting point	High	Lower
Texture / structure	Brittle	—
Conducts electricity when dissolved	Yes, well	Poorly

Covalent bonds are further split into polar and nonpolar. If one atom hogs the shared electrons, the bond is polar covalent; if the atoms share electrons equally (or nearly so), it's nonpolar covalent. This is judged by the difference in electronegativity — since the AP exam doesn't provide an electronegativity chart, use periodic table proximity instead: atoms close to each other (e.g., selenium and iodine) form more nonpolar bonds, while atoms farther apart (e.g., selenium and oxygen) form more polar bonds.

2. Bond Energy and the Potential Energy Curve

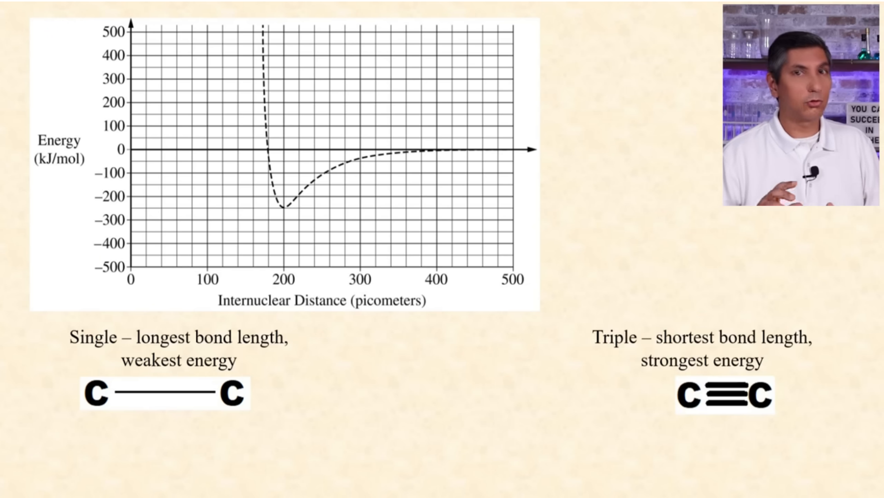
As two atoms move farther apart, their potential energy rises; as they approach and form a bond, potential energy drops. The curve below plots this relationship — energy (kJ/mol) against internuclear distance (picometers):



Potential energy vs. internuclear distance curve.

The bond length is the internuclear distance at the lowest point of the curve — 200 picometers in this example. The bond energy is the absolute value of the potential energy at that same point — 250 kJ/mol here.

Bond order (single, double, or triple) affects both length and strength: single bonds are the longest and weakest, triple bonds are the shortest and strongest, and double bonds fall in between.



Single – longest bond length, weakest energy

$C-C$

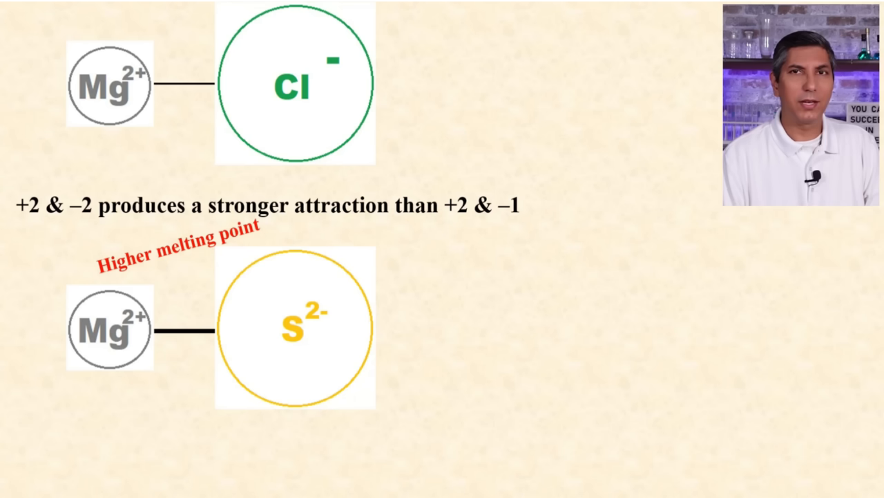
Triple – shortest bond length, strongest energy

$C\equiv C$

Bond order comparison: single bonds are longer/weaker; triple bonds are shorter/stronger.

3. Coulomb's Law and Ionic Melting Points

The greater the magnitude of ionic charge, the stronger the attraction between ions. Magnesium chloride (Mg^{2}■ / Cl■) has a fairly strong attraction, but magnesium sulfide (Mg^{2}■ / S^{2}■) is even more strongly attracted because sulfide carries a –2 charge — giving magnesium sulfide a higher melting point than magnesium chloride.



+2 & –2 produces a stronger attraction than +2 & –1

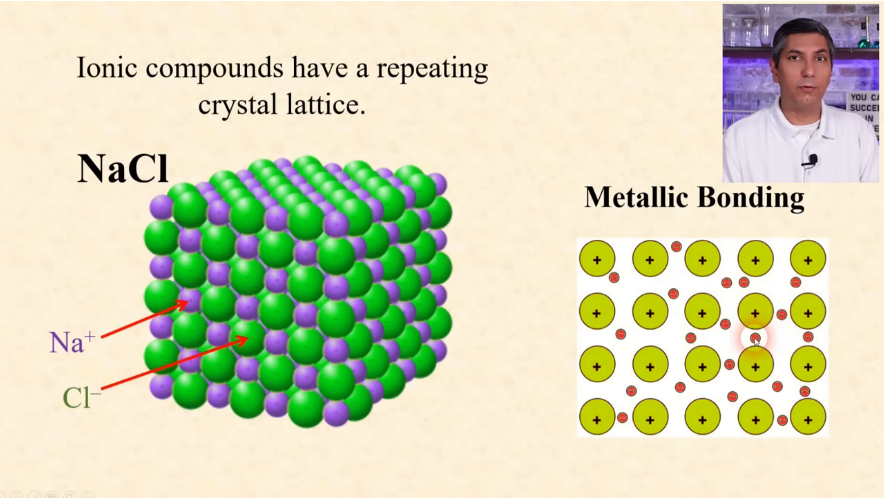
Higher melting point

Charge magnitude comparison: a +2/–2 pairing attracts more strongly than +2/–1.

Ion size matters too: larger ions attract neighboring ions more weakly, since their charge is spread farther from the surface. Magnesium oxide (O^{2}■ is smaller than S^{2}■) has an even higher melting point than magnesium sulfide. Rule of thumb: compare charge first; if it's a tie, compare ionic size.

4. Ionic Crystal Lattices

Ionic compounds like sodium chloride are not made of independent, freely floating NaCl units. Instead, they form a repeating, three-dimensional crystal lattice of alternating cations and anions.



Ionic compounds have a repeating crystal lattice.

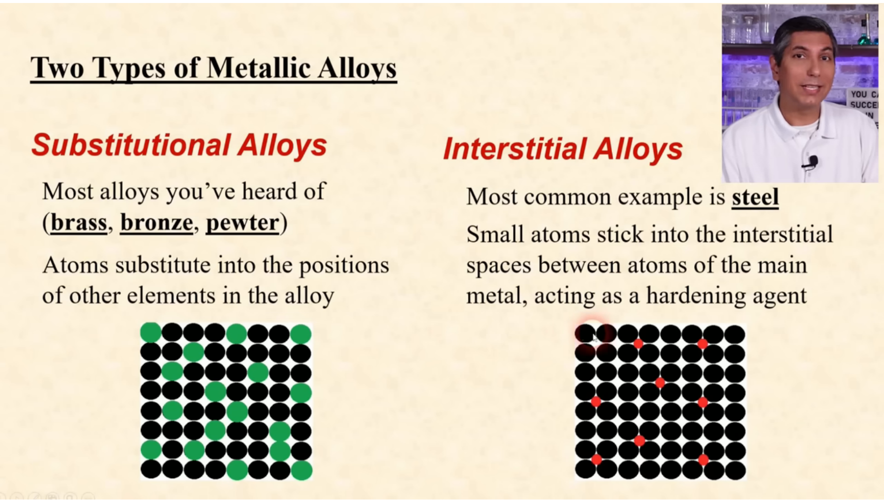
NaCl

Metallic Bonding

Repeating crystal lattice structure of sodium chloride (NaCl).

5. Metallic Bonding and Alloys

Metals form positively charged cations surrounded by a “sea” of delocalized valence electrons that can move freely — this explains why metals conduct electricity so well. When other elements combine with metals, they form alloys, which come in two main types:



Two Types of Metallic Alloys

Substitutional Alloys

Most alloys you’ve heard of (**brass**, **bronze**, **pewter**)

Atoms substitute into the positions of other elements in the alloy

Interstitial Alloys

Most common example is **steel**

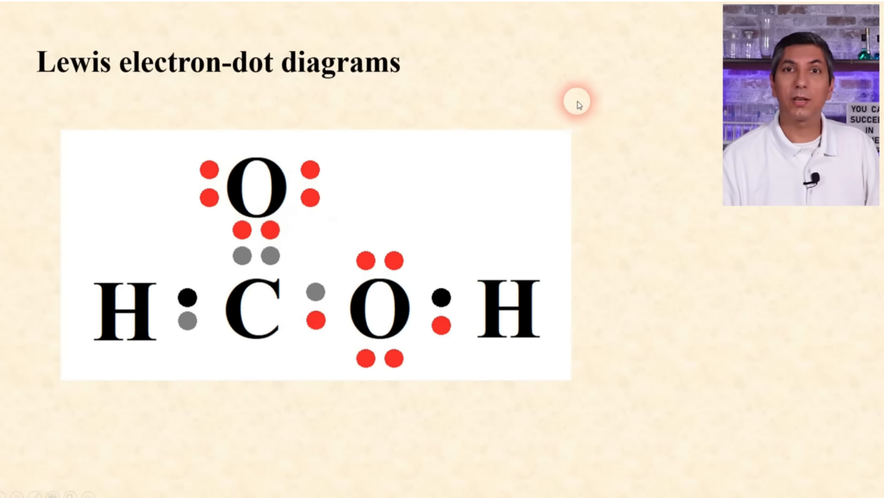
Small atoms stick into the interstitial spaces between atoms of the main metal, acting as a hardening agent

Substitutional alloys (e.g., brass, bronze, pewter) vs. interstitial alloys (e.g., steel).

In substitutional alloys, atoms of one element take the place of some atoms of the primary metal — in brass, zinc atoms substitute into positions normally held by copper atoms. In interstitial alloys, smaller atoms squeeze into the gaps between the primary metal's atoms; in steel, small carbon atoms occupy the interstitial spaces between iron atoms, acting as a hardening agent.

6. Lewis Electron-Dot Diagrams

When drawing a Lewis structure, start from the outside atoms and work inward. Hydrogen is stable with 2 valence electrons; most other atoms aim for a full octet (8). If a central atom ends up short of an octet, move a lone pair from a neighboring atom to form a double bond.



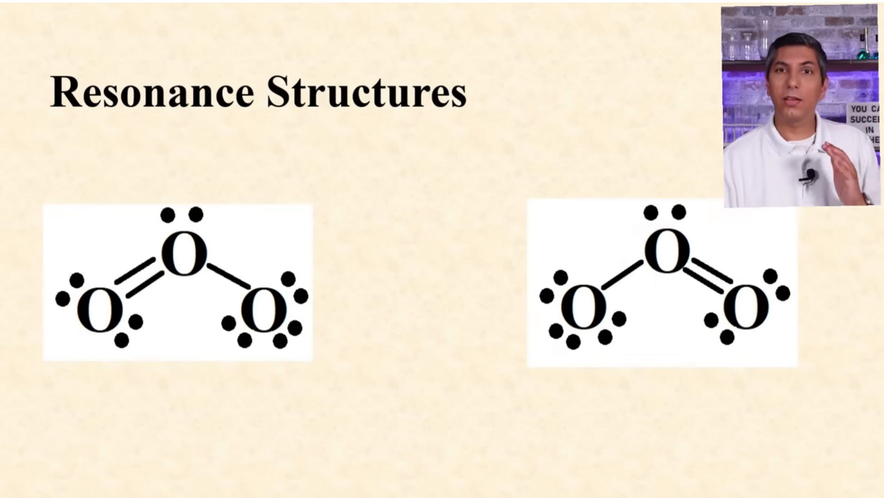
Lewis electron-dot diagrams

Completed Lewis electron-dot structure, with a double bond added to satisfy the central atom's octet.

Some molecules need more than 8 valence electrons on the central atom — this is called an expanded octet (e.g., xenon tetrafluoride, which places 8 dots directly around the central atom).

7. Resonance Structures

Sometimes more than one valid Lewis structure can be drawn for the same molecule — these are called resonance structures. Ozone (O■) is the classic example: the double bond can be drawn on either side of the central oxygen, and both structures are equally acceptable.

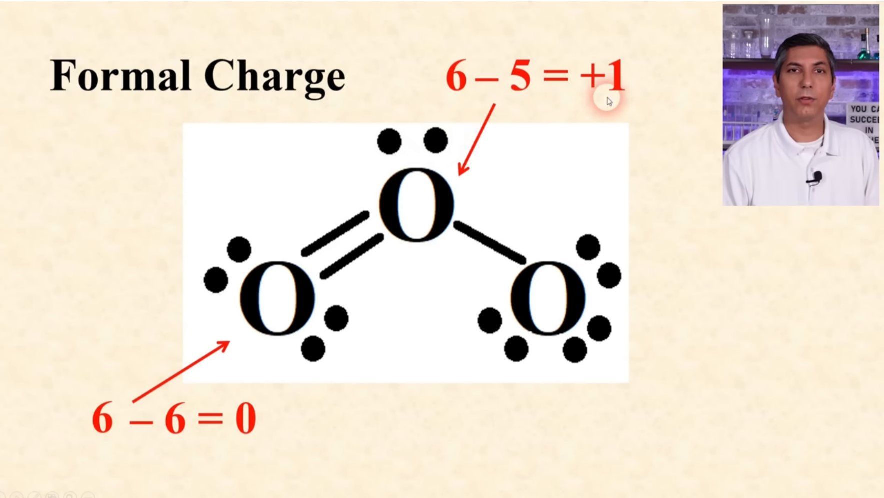


Resonance Structures

Resonance structures of ozone (O■) — the double bond position shifts between structures.

8. Formal Charge

Formal charge = (valence electrons the atom normally has) – (electrons assigned to it in the Lewis diagram, counting each bond as one electron). In ozone, this works out to 0, +1, and –1 across the three oxygens — and the formal charges always sum to the molecule's overall charge.



Formal charge calculation shown on ozone's resonance structure.

In a neutral molecule, the most stable Lewis structure is usually — though not always — the one where every atom has a formal charge of zero.

9. VSEPR, Hybridization, and Molecular Geometry

Every single bond is one sigma (σ) bond; every double bond is one sigma + one pi (π) bond; every triple bond is one sigma + two pi bonds. Hybridization is found by adding the number of atoms a central atom touches to its number of unshared (lone) electron pairs: 2 total → sp, 3 total → sp^2, 4 total → sp^3 (and 5 or 6 total extend to sp^3d and sp^3d^2).

The table below — rebuilt from the video's summary chart — lists molecular geometry and bond angle for every combination of bonded atoms and lone pairs on a central atom:

Atoms Touching	Unshared Pairs	Hybridization	Molecular Geometry	Bond Angle
4	0	sp3	Tetrahedral	109.5°
3	1	sp3	Trigonal Pyramidal	107°
2	2	sp3	Bent	104.5°
3	0	sp2	Trigonal Planar	120°
2	1	sp2	Bent / Angular	~120°
2	0	sp	Linear	180°
6	0	sp3d2	Octahedral	90°
5	1	sp3d2	Square Pyramidal	~90°
4	2	sp3d2	Square Planar	90°
3	3	sp3d2	T-shaped	~90°

Atoms Touching	Unshared Pairs	Hybridization	Molecular Geometry	Bond Angle
5	0	sp3d	Trigonal Bipyramidal	90° & 120°
4	1	sp3d	Seesaw	~90° & 120°
3	2	sp3d	T-shaped	~90°

The AP exam is generally lenient on exact bond angles — the values you're expected to know are 109.5°, 120°, 90°, and 180°.

Summary: Unit 2 covers how atoms bond (ionic vs. covalent, polar vs. nonpolar), how bond energy and order relate to atomic distance, how ionic charge and size drive melting points via Coulomb's law, how metals and alloys are structured, and how to draw and analyze Lewis structures — including resonance, formal charge, and the VSEPR-based molecular geometries.