

Unit 2: Molecular and Ionic Compound Structure and Properties

Unit 2 explains how atoms stick together and why the way they stick together determines molecular shape and the properties of solids. It covers the three types of chemical bonds, what bond strength depends on, how ionic lattices and metal alloys are organized, how to draw Lewis structures, and how VSEPR predicts molecular geometry.

AP Chemistry

Molecular and Ionic Compound Structure and Properties

About 12 minutes to read

How to use this guide

Read it in order the first time because the topics build on each other. Bond types set up the potential energy picture, the energy picture explains bond strength, bond strength explains why lattices and alloys hold together the way they do, and Lewis structures plus VSEPR turn all of that into molecular shapes you can predict on paper.

After the first read, use the trap boxes to review the distinctions that exam questions test most often. Finish with the practice questions, then complete the recall check on the last page out loud and note any items you cannot explain yet.

What this unit is worth. Molecular and Ionic Compound Structure and Properties is about 7 to 9 percent of the AP Chemistry exam. The ideas also carry into later units, because intermolecular forces in Unit 3 build directly on the polarity you learn here, and reaction energetics in Unit 6 depends on bond energies.

2.1 Types of Chemical Bonds

Chemical bonds come in three types, and the type depends on what the atoms do with their electrons. An **ionic bond** forms between a metal and a nonmetal, where electrons are transferred. Sodium chloride is the standard example. Sodium gives up an electron and chlorine takes it, producing Na^+ and Cl^- . But bonding is a continuum, not two separate boxes. A very polar covalent bond behaves almost as if electrons were fully transferred, and even a mostly ionic bond keeps some covalent character. Classify by the atoms involved, but do not treat the boundary as a wall.

A **covalent bond** shares electrons. When the two atoms have similar electronegativity, the sharing is nearly equal and the bond is **nonpolar covalent**. H_2 and Cl_2 are the cleanest examples. When the atoms have different electronegativities, the sharing is unequal and the bond is **polar covalent**. In HCl, chlorine (electronegativity 3.16) pulls harder than hydrogen (2.20), so the bond carries partial charges: δ^+ on hydrogen, δ^- on chlorine. That charge separation is the **bond dipole**, and it grows as the electronegativity difference grows.

Metallic bonding is the third type. In a metal, valence electrons are delocalized over an array of positive metal ions rather than belonging to individual atoms. That delocalization is what lets metals conduct electricity and bend without shattering.

Trap. Students see metal plus nonmetal and treat ionic as a completely different category from covalent. The bond types form a continuum. If a question asks you to classify, answer from the atoms involved, but never claim a bond is 100 percent ionic or 100 percent covalent.

2.2 Intramolecular Force and Potential Energy

The **potential energy vs. distance graph** shows what happens as two atoms approach each other. Far apart, attraction is weak. As they approach, attraction lowers the potential energy until it reaches a minimum. That minimum sits at the **equilibrium bond length**, the most stable separation between the two bonded atoms. Push the atoms closer than that and repulsion drives the energy back up steeply.

The depth of the well equals the **bond energy**, the energy required to pull the two atoms completely apart. A deeper well means a stronger bond. **Bond order**, the number of shared electron pairs, connects to this directly. A double bond is shorter and stronger than a single bond between the same two atoms, and a triple bond is shorter and stronger still. Carbon-carbon bonds show the pattern: C–C is 154 pm, C=C is 134 pm, and C $\equiv$ C is 120 pm. The length drops as the order rises.

Trap. Breaking a bond requires energy input. Students sometimes say bond breaking releases energy because they are thinking of combustion, where the net release comes from forming stronger new bonds. Pulling two bonded atoms apart always costs energy. The bond energy is the price of separation.

For ionic bonds, strength follows **Coulomb's law**. The attraction is proportional to the product of the ion charges and inversely proportional to the square of the distance between them, so larger charges and smaller ions make stronger ionic bonds. MgO melts at about 2852 °C while NaCl melts at 801 °C, because Mg^{2+} and O^{2-} carry twice the charge of Na^{+} and Cl^{-} and are smaller ions. That is Coulomb's law showing up in a physical property you can measure.

2.3 Structure of Ionic Solids

An **ionic crystal lattice** is a three-dimensional array of ions arranged to maximize attraction between opposite charges and minimize repulsion between like charges, exactly what Coulomb's law predicts. In table salt, each Na^{+} is surrounded by six Cl^{-} ions, and each Cl^{-} by six Na^{+} ions. The lattice explains the characteristic properties. Ionic solids are brittle because shifting a layer lines up like charges, which repel and crack the crystal. They conduct electricity only when molten or dissolved, because only then are the ions free to move.

2.4 Structure of Metals and Alloys

The **sea of electrons model** pictures a metal as an ordered array of positive metal ions surrounded by delocalized valence electrons. The mobile electrons explain conductivity, malleability, and ductility. The ions can slide past each other without breaking the bonding because the electron sea flows along with them.

An **alloy** mixes a second element into the metal, and the two types work differently. In an **interstitial alloy**, smaller atoms occupy the spaces between the larger metal atoms. Carbon in iron makes steel. The small atoms distort the lattice, which makes it harder for layers to slide, so the alloy is harder and more rigid than the pure metal. In a **substitutional alloy**, atoms of a similar size replace some of the metal atoms in the lattice. Zinc replacing copper makes brass. The size match means less distortion, but the different atoms still disrupt the regular sliding of layers.

Trap. Students mix up which alloy is which. Interstitial means the added atoms sit in the gaps between metal atoms, so they must be much smaller. Substitutional means they take the place of metal atoms, so they must be about the same size. Steel means small carbon atoms squeezed between iron atoms. Brass means zinc atoms swapping in for copper atoms.

2.5 Lewis Diagrams

A **Lewis diagram** shows the valence electrons and the bonds in a molecule or ion, built by a fixed set of rules. Count the total valence electrons, adjusting for the charge on ions. Connect the atoms with single bonds. Distribute the remaining electrons as lone pairs to satisfy octets. Convert lone pairs to multiple bonds where needed. For H_2O , oxygen brings 6 valence electrons and each hydrogen brings 1, for 8 total. Two O–H single bonds use 4 electrons, the remaining 4 go on oxygen as two lone pairs, and oxygen has its octet.

The **octet rule** is the guideline that atoms ideally end up with eight valence electrons. It has exceptions, and the exam tests them. Species with an odd number of valence electrons cannot satisfy it. NO has 11 valence electrons, so one atom will always fall short. Some atoms end up with fewer than eight. Boron in BF_3 has only 6. Atoms in period 3 and below can expand past eight. Phosphorus in PCl_5 has 10. Do not force an octet where the electron count forbids it.

2.6 Resonance and Formal Charge

Resonance applies when more than one valid Lewis structure can be drawn for the same arrangement of atoms. Ozone (O_3) is the classic case. You can draw the double bond on the left oxygen or the right oxygen, and both structures are valid. The real molecule is neither. It is a hybrid, an average of the contributing forms, with two identical bonds that each fall between a single and a double bond in length and strength.

Formal charge is the bookkeeping tool for choosing among nonequivalent Lewis structures. For each atom, the formal charge equals its valence electrons minus the electrons assigned to it: all of its lone-pair electrons plus half of its bonding electrons. The best structure minimizes formal charges, and any negative formal charge should sit on the most electronegative atom. For CO_2 , the structure $\text{O}=\text{C}=\text{O}$ gives every atom a formal charge of zero, which is why it wins over alternatives with charges.

Trap. Two mistakes show up constantly. First, students think resonance structures are real molecules flipping back and forth. They are not. The contributing structures are drawings. The hybrid is the reality. Second, students treat formal charge as the actual charge on an atom. It is a bookkeeping value for comparing structures, not a measured physical quantity.

2.7 Molecular Geometry and VSEPR

VSEPR theory says electron domains around a central atom arrange themselves to minimize repulsion, and that arrangement predicts molecular geometry. Count the domains, meaning bonding pairs plus lone pairs. Find the arrangement that spreads them out. Then name the shape using only the atoms.

DOMAINS	ELECTRON GEOMETRY	MOLECULAR SHAPE (EXAMPLES)	BOND ANGLE
2	Linear	Linear: CO_2 , BeCl_2	180°
3	Trigonal planar	Trigonal planar: BF_3	120°
4	Tetrahedral	Tetrahedral: CH_4 ; trigonal pyramidal: NH_3 ; bent: H_2O	109.5° (less with lone pairs)

With four domains, the molecular shape depends on the lone pairs. CH_4 has none and is tetrahedral at 109.5° . NH_3 has one lone pair and is trigonal pyramidal at about 107° . H_2O has two lone pairs and is bent at about 104.5° . The **bond angle** shrinks as lone pairs are added because lone pairs repel more strongly than bonding pairs.

Whether a molecule is polar depends on geometry, not just bond polarity. A molecule has a **molecular dipole moment**, meaning it is polar, only if it has polar bonds AND their dipoles do not cancel through symmetry. CO_2 has two polar $\text{C}=\text{O}$ bonds, but they point in exactly opposite directions in the linear molecule, so the dipoles cancel and CO_2 is nonpolar. H_2O has polar $\text{O}-\text{H}$ bonds in a bent shape, so the dipoles add up and water is polar. CCl_4 is tetrahedral and symmetric, so it is nonpolar despite four polar bonds. CH_3Cl has the same tetrahedral electron geometry, but the atoms differ, so the dipoles do not cancel and it is polar.

Trap. Polar bonds do not guarantee a polar molecule. This is one of the most tested distinctions in the unit. Always check the geometry before declaring polarity. If the molecule is symmetric, the dipoles cancel.

2.7, continued: Hybridization, Sigma and Pi Bonds

Hybridization describes the mixing of atomic orbitals to match the observed geometry around an atom. Four electron domains (tetrahedral) correspond to sp^3 hybridization, as in CH_4 . Three domains (trigonal planar) correspond to sp^2 , as in BF_3 . Two domains (linear) correspond to sp , as in CO_2 and $BeCl_2$. The bond angles follow the geometry: sp^3 is 109.5° , sp^2 is 120° , sp is 180° .

Every covalent bond contains one **sigma (σ) bond**, formed by head-on orbital overlap. It is the strongest component and the only bond in a single bond. A **pi (π) bond** forms by sideways overlap of p orbitals. It is weaker than a sigma bond, and it prevents rotation about the bond axis. The count follows directly. A single bond is one sigma. A double bond is one sigma plus one pi. A triple bond is one sigma plus two pi bonds.

That locked rotation is what makes **geometric isomers** possible. In 2-butene, the double bond holds the two ends fixed, so the two CH_3 groups can sit on the same side (cis-2-butene) or on opposite sides (trans-2-butene). Same formula, different spatial arrangement, different properties. Without a pi bond there is no locked rotation, and without locked rotation there are no geometric isomers.

BOND TYPE	COMPOSITION	ROTATION
Single	1 sigma	Free
Double	1 sigma + 1 pi	Locked
Triple	1 sigma + 2 pi	Locked

Practice Questions

Original questions written for this guide in the style of the AP exam. Answers and explanations are on the next page, so complete the questions before checking them.

1. The bond between magnesium and chlorine in MgCl_2 is best classified as

- A. nonpolar covalent
- B. polar covalent
- C. ionic
- D. metallic

2. Which compound has the strongest ionic bonding?

- A. NaCl
- B. MgO
- C. KCl
- D. NaBr

3. The molecular geometry of NH_3 is

- A. trigonal planar
- B. tetrahedral
- C. trigonal pyramidal
- D. bent

4. Which molecule is polar?

- A. CO_2
- B. CCl_4
- C. H_2O
- D. BF_3

Answer Key

1. C. Magnesium is a metal and chlorine is a nonmetal, so electrons transfer and the bond is ionic. A and B describe covalent sharing, which does not fit a metal-nonmetal pair. D is wrong because metallic bonding involves delocalized electrons among metal atoms, not a metal-nonmetal compound.

2. B. Coulomb's law says ionic attraction grows with the product of the charges and shrinks with distance. Mg^{2+} and O^{2-} give the largest charge product of the four choices, and both ions are small. A, C, and D all involve +1 and -1 ions, so their attraction is weaker.

3. C. Nitrogen has four electron domains: three N-H bonds and one lone pair. Four domains give tetrahedral electron geometry, but the molecular shape counts only atoms, so with one lone pair the shape is trigonal pyramidal. A ignores the lone pair. B names the electron geometry rather than the molecular shape. D is the shape for two bonds plus two lone pairs, as in H_2O .

4. C. H_2O has polar O-H bonds and a bent geometry, so the dipoles add up instead of canceling. A is linear, so its two bond dipoles cancel exactly. B is tetrahedral and symmetric, so its four dipoles cancel. D is trigonal planar and symmetric, so its three dipoles cancel.

One-Page Recall Check

- Describe the three bond types and what happens to electrons in each.
- Explain why bonding is a continuum rather than two separate categories.
- State how a bond dipole relates to electronegativity difference, and assign δ^+ and δ^- in HCl.
- Sketch the potential energy vs. distance graph and label the equilibrium bond length and the bond energy.
- Explain how bond order relates to bond length and bond strength, with the C–C, C=C, C $\equiv$ C numbers.
- State Coulomb's law and use it to explain why MgO melts far above NaCl.
- Describe the ionic crystal lattice and explain why ionic solids are brittle and conduct only when molten or dissolved.
- Explain the sea of electrons model and what it accounts for in metals.
- Distinguish interstitial from substitutional alloys, with steel and brass as examples.
- Draw the Lewis diagram of H₂O from scratch using the standard rules.
- Name three exceptions to the octet rule with an example of each.
- Explain what resonance means for ozone, and why the contributing structures are not real molecules.
- Define formal charge and state the rule for picking the best Lewis structure.
- Use VSEPR to predict the shape and bond angle of CH₄, NH₃, H₂O, BF₃, and CO₂.
- Explain why CO₂ is nonpolar while H₂O is polar, even though both have polar bonds.
- Match sp, sp², and sp³ hybridization to their geometries and bond angles.
- State the sigma and pi composition of single, double, and triple bonds.
- Explain how geometric isomers arise, using cis- and trans-2-butene.

Study this unit in Rycal. Open the AP Chemistry deck at <https://rycal.web.app/apchem> and drill the Unit 2 cards. The deck covers the terms in this guide, and its practice questions target the same traps named here. If you have a test date, add it in the Test Planner. You can also start your next review with a Brain Dump, then check what you missed against this guide.

Key terms for this unit

Nonpolar covalent bond, Polar covalent bond, Bond dipole (partial charges), Ionic bond, Metallic bonding, Potential energy vs. distance graph, Equilibrium bond length, Bond energy, Bond order, Coulomb's law and ionic bond strength, Ionic crystal lattice, Sea of electrons model, Interstitial alloy, Substitutional alloy, Lewis diagram (Lewis structure), Resonance, Formal charge, Octet rule, VSEPR theory, Molecular geometries, Bond angle, Molecular dipole moment, Hybridization (hybrid orbitals), Sigma (σ) bond, Pi (π) bond, Geometric isomers

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