

Unit 1: Atomic Structure and Properties

Unit 1 is the counting unit. It covers the mole and molar mass, mass spectroscopy and isotopes, elemental composition of pure substances and mixtures, electron configurations, photoelectron spectroscopy, and the periodic trends that all follow from atomic structure.

AP Chemistry

Atomic 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. The mole lets you count particles, mass spectroscopy explains where atomic masses come from, composition work applies the mole to real samples, and then the unit shifts to the atom itself: electron configurations, the evidence from PES, and the periodic trends that Coulomb's law predicts. Exam questions in this unit are mostly calculations and data interpretation, so work the examples with a pencil rather than just reading them.

After the first read, use the trap boxes for the distinctions the exam tests most often, especially empirical versus molecular formulas and the periodic trend directions. Finish with the practice questions, then do the recall check on the last page out loud and note anything you cannot explain yet.

What this unit is worth. Atomic Structure and Properties is about 7 to 9 percent of the AP Chemistry exam. The mole, dimensional analysis, and electron configurations show up in every unit that follows.

1.1 Moles and Molar Mass

A **mole** is the chemist's counting unit. One mole of anything contains 6.022×10^{23} particles, a number called **Avogadro's number**. The particles being counted depend on the substance. For a molecular substance like water, one mole contains 6.022×10^{23} molecules. For an ionic compound like NaCl, there are no molecules, so the countable unit is the **formula unit**, the lowest whole-number ratio of ions. One mole of NaCl contains 6.022×10^{23} formula units of NaCl.

The **molar mass** of a substance is the mass of one mole of it, in grams per mole. It is numerically equal to the average atomic or formula mass in **atomic mass units (amu)**. Water has an average molecular mass of about 18.02 amu, so its molar mass is 18.02 g/mol. This gives you the bridge between the mass you measure on a balance and the number of particles you are actually working with.

Dimensional analysis is the method for moving between grams, moles, and particles. You multiply by conversion factors and cancel units. For example, convert 36.0 g of H₂O to molecules. First grams to moles: $36.0 \text{ g} \times (1 \text{ mol} / 18.02 \text{ g}) = 2.00 \text{ mol}$. Then moles to molecules: $2.00 \text{ mol} \times (6.022 \times 10^{23} \text{ molecules} / 1 \text{ mol}) = 1.20 \times 10^{24} \text{ molecules}$. Write every unit at every step so you can see what cancels.

Trap. Molar mass converts grams to moles, and Avogadro's number converts moles to particles. Students often multiply grams directly by Avogadro's number and skip the mole step. The units will not cancel if you do that, which is exactly what dimensional analysis is there to catch.

1.2 Mass Spectroscopy of Elements

A **mass spectrum** is a plot of ion abundance versus mass for a sample. For a single element, each peak corresponds to an **isotope**, which is an atom of the element with a different number of neutrons and therefore a different mass. The height of each peak shows the **relative abundance** of that isotope, meaning the percentage of the element's atoms in nature that have that mass.

The atomic mass on the periodic table is the **average atomic mass**, the weighted average of the isotope masses using their relative abundances as weights. Chlorine is the standard example. About 75.78% of chlorine atoms are Cl-35 with a mass of 34.969 amu, and about 24.22% are Cl-37 with a mass of 36.966 amu. The average atomic mass is $(0.7578 \times 34.969) + (0.2422 \times 36.966) = 26.50 + 8.95 = 35.45$ amu, which is the value printed on the periodic table.

Notice that the average is closer to 35 than to 37 because the lighter isotope is much more abundant. The weighted average always leans toward the isotope with the greater relative abundance. On the exam, convert each percent abundance to a decimal before multiplying, and check that your answer falls between the lightest and heaviest isotope masses.

Trap. Average atomic mass is not the simple average of the isotope masses. A plain average of 34.969 and 36.966 gives 35.97, which is wrong. The abundances are the weights, so the 75.78% isotope dominates and pulls the answer to 35.45.

1.3 Elemental Composition of Pure Substances

The **law of definite proportions** says a pure compound always contains its elements in the same fixed ratio by mass. Every sample of water is 11.2% hydrogen and 88.8% oxygen by mass, no matter where the water came from. This is what makes percent composition a reliable tool for identifying compounds.

The **empirical formula** gives the lowest whole-number ratio of atoms in a compound, while the **molecular formula** gives the actual number of atoms per molecule. To find an empirical formula from percent composition, convert each percent to grams (assume a 100 g sample), convert grams to moles, then divide all mole values by the smallest one. A compound that is 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen gives 3.33 mol C, 6.65 mol H, and 3.33 mol O per 100 g. Dividing by 3.33 gives a 1:2:1 ratio, so the empirical formula is CH₂O.

If you also know the molar mass, you can get the molecular formula. CH₂O has an empirical formula mass of 30.03 g/mol. If the compound's molar mass is 180 g/mol, the ratio $180 / 30.03$ is about 6, so the molecular formula is C₆H₁₂O₆. The molecular formula is always a whole-number multiple of the empirical formula.

Trap. Dividing by the smallest mole value gives the ratio, but the numbers must come out close to whole numbers. If you get 1.5, multiply everything by 2 rather than rounding to 2. Rounding a ratio like 1:1.5:1 to 1:2:1 changes the compound.

1.4 Composition of Mixtures

A **pure substance** contains only one type of atom, molecule, or formula unit. A **mixture** contains two or more types of particles whose proportions can vary, and it can be homogeneous (uniform throughout, like salt water) or heterogeneous (visibly nonuniform, like sand and water). The law of definite proportions applies to pure compounds, not to mixtures, because a mixture has no fixed composition.

Elemental analysis is the experimental determination of how many atoms of each element a sample contains, and it is used to identify a substance's composition and check its purity. For a mixture, the **elemental composition** can be worked out from the masses of the components. If a 10.0 g sample of a copper-zinc mixture contains 6.5 g of copper, the mixture is 65.0% copper and 35.0% zinc by mass. The same idea works in reverse. If you know a mixture is 40.0% sodium chloride by mass and you have 25.0 g of the mixture, it contains $0.400 \times 25.0 = 10.0$ g of NaCl.

Trap. Percent composition questions about mixtures look like empirical formula questions but are not. An empirical formula assumes a pure compound with a fixed ratio. A mixture's composition varies, so you work from the given masses directly instead of converting to a formula.

1.5 Atomic Structure and Electron Configuration

Atoms are built from three **subatomic particles**. Protons carry a +1 charge and neutrons are neutral, and both sit in the nucleus. Electrons carry a -1 charge and occupy shells around the nucleus. The number of protons defines the element. In a neutral atom, the proton count equals the electron count.

An **electron configuration** describes how the electrons are arranged in shells and subshells, and it is written by filling the lowest-energy subshells first, which is the **aufbau principle**. Sodium, with 11 electrons, is $1s^2 2s^2 2p^6 3s^1$. The **valence electrons** are the ones in the outermost shell. For sodium that is the single 3s electron, and it is the valence electrons that determine how an element behaves chemically. The inner electrons are **core electrons**.

What holds electrons to the atom is electrostatic attraction, described by **Coulomb's law**. The force between charges is proportional to $q_1 q_2 / r^2$, so a larger nuclear charge pulls harder and a shorter electron-to-nucleus distance pulls harder too. Inner electrons partially block the nucleus from outer electrons, an effect called **shielding**. The net positive charge an electron actually feels after shielding is the **effective nuclear charge**.

The **ionization energy** is the energy needed to remove an electron from a gaseous atom or ion, and Coulomb's law explains its patterns. An electron that is closer to the nucleus, or that feels a larger effective nuclear charge, is held more tightly and costs more energy to remove. This is why the first ionization energy is always smaller than the second. Removing the first electron from sodium takes its lone 3s valence electron, but removing a second means pulling a core electron out of the stable $n = 2$ shell, much closer to the nucleus.

Trap. Ionization energy explanations need both parts of Coulomb's law: charge and distance. Saying "the nuclear charge is larger" is only half the reasoning. Name the distance too, as in "the electron is farther from the nucleus, so the attraction is weaker despite the larger charge."

1.6 Photoelectron Spectroscopy

Photoelectron spectroscopy (PES) gives experimental evidence for the shell model. The technique shines high-energy light on atoms and measures the **binding energy** of the ejected electrons, which is the energy needed to remove an electron from its subshell. Each peak in a PES spectrum corresponds to a subshell. The peak's position on the energy axis tells you the binding energy, and the peak's height tells you how many electrons are in that subshell.

Read a PES spectrum from right to left, highest binding energy to lowest. The rightmost peak is always the 1s subshell, because those electrons are closest to the nucleus and hardest to remove. Moving left, each peak is a subshell farther out with a lower binding energy. A spectrum with three peaks at 19.3, 1.36, and 0.80 MJ/mol, with heights in the ratio 2:2:1, reads as $1s^2 2s^2 2p^1$, which is boron. The peak heights give the electron counts directly, so you can reconstruct the configuration without memorizing filling order.

PES also explains the big jumps in successive ionization energies from topic 1.5. The large gap between the 2s and 1s peaks shows why removing a core electron costs so much more than removing a valence electron. The spectrum is the measurement behind the claim.

Trap. Higher binding energy means the electron is held more tightly, which means it is closer to the nucleus. Students sometimes read the tallest peak or the leftmost peak as the innermost electrons. Position on the energy axis is what matters, and the innermost subshell sits at the highest energy, on the right.

1.7 Periodic Trends

Periodicity is the repeating pattern of properties across periods and down groups, and it happens because electron configurations repeat. Elements in the same group have the same number of valence electrons, so they behave similarly. Everything in this topic is explained by the **shell model** plus Coulomb's law and shielding.

The **atomic radius trend** runs two directions. Across a period, radius decreases because the effective nuclear charge rises while electrons are added to the same shell, so the stronger pull draws the electron cloud in. Down a group, radius increases because each new period adds a shell, and the outermost electrons sit farther from the nucleus despite the larger charge.

Ionic radius follows from what forming an ion does. Cations are smaller than their parent atoms because they lose their valence shell entirely. Na^+ is much smaller than Na because the $n = 3$ shell is gone. Anions are larger than their parent atoms because the added electron increases repulsion among the valence electrons while the nuclear charge stays the same.

Electronegativity measures how strongly an atom attracts shared electrons in a bond. It increases across a period and decreases down a group, for the same Coulomb's law reasons as atomic radius. Fluorine sits at the top right and is the most electronegative element. **Electron affinity**, the energy change when an atom gains an electron, follows a similar logic and can also be explained with Coulomb's law, though its trend has more exceptions and the exam tests it less often.

Trap. Every trend in this topic needs a Coulomb's law justification, not just the direction. "Atomic radius decreases across a period" earns partial credit at best. The full answer names the mechanism: effective nuclear charge increases while the shell number stays the same, so the attraction per electron strengthens and the radius shrinks.

1.8 Valence Electrons and Ionic Compounds

Main-group elements form ions by gaining or losing electrons to reach a noble-gas configuration, and the charge follows from the valence electron count. An element with one valence electron, like sodium, loses it to become Na^+ . An element with seven valence electrons, like chlorine, gains one to become Cl^- .

Predicting ionic charge is a matter of counting how many electrons must move to empty or fill the outer shell.

Because group members have the same number of valence electrons, elements in the same group form **analogous compounds** with similar formulas. Magnesium and calcium are both group 2, so just as magnesium forms MgCl_2 , calcium forms CaCl_2 . If you know one compound in a group, you can predict the formulas of the rest.

Confusions That Cost Points

PAIR	HOW TO KEEP THEM STRAIGHT
Mole vs molar mass	The mole is a count: 6.022×10^{23} particles. Molar mass is a mass: grams per one mole. Avogadro's number converts moles to particles, molar mass converts moles to grams.
Empirical vs molecular formula	Empirical is the simplest whole-number ratio. Molecular is the actual atom count per molecule, always a whole-number multiple of the empirical formula. You need the molar mass to go from one to the other.
Pure substance vs mixture	A pure substance has a fixed composition that the law of definite proportions describes. A mixture's proportions vary, so percent composition is computed from the given masses, not from a formula.
Ionization energy vs binding energy	Ionization energy is the energy to remove a specific electron, usually the outermost. Binding energy from PES is the same idea measured per subshell, which is why PES shows the jumps between shells directly.
Cation vs anion radius	Cations are smaller than their atoms because a whole shell is lost. Anions are larger because added electron repulsion spreads the same shell wider.
Electronegativity vs electron affinity	Electronegativity is attraction for shared electrons within a bond. Electron affinity is the energy change when a free atom gains an electron. Related ideas, different definitions.

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. A sample of CO_2 has a mass of 22.0 g. How many molecules does it contain?
 - A. 8.30×10^{-25}
 - B. 3.01×10^{23}
 - C. 6.02×10^{23}
 - D. 1.32×10^{25}
2. Element X has two isotopes: X-63 with a mass of 62.93 amu and 69.2% abundance, and X-65 with a mass of 64.93 amu and 30.8% abundance. What is the average atomic mass of X?
 - A. 63.93 amu
 - B. 63.55 amu
 - C. 64.93 amu
 - D. 63.00 amu
3. The PES spectrum of an element shows three peaks at 19.3, 1.36, and 0.80 MJ/mol with heights in the ratio 2:2:1. Which element is it?
 - A. Lithium
 - B. Beryllium
 - C. Boron
 - D. Carbon
4. Which of the following best explains why atomic radius decreases across a period?
 - A. Electrons are added to a new shell, which shields the nucleus more effectively
 - B. Effective nuclear charge increases while electrons are added to the same shell, strengthening the attraction
 - C. Ionization energy decreases, allowing the electron cloud to contract
 - D. Protons are added to the nucleus, increasing repulsion between electrons

Answer Key

1. B. Molar mass of $\text{CO}_2 = 12.01 + 2(16.00) = 44.01 \text{ g/mol}$. Moles = $22.0 \text{ g} / 44.01 \text{ g/mol} = 0.500 \text{ mol}$.

Molecules = $0.500 \text{ mol} \times 6.022 \times 10^{23} = 3.01 \times 10^{23}$. A comes from dividing by Avogadro's number instead of multiplying. C is the number of molecules in a full mole, forgetting to convert 22.0 g to 0.500 mol first. D multiplies by Avogadro's number without converting grams to moles.

2. B. Average = $(0.692 \times 62.93) + (0.308 \times 64.93) = 43.55 + 20.00 = 63.55 \text{ amu}$. A averages the two masses without weighting by abundance. C reports the heavier isotope's mass instead of the average. D rounds the lighter isotope's mass and ignores the second isotope entirely.

3. C. The rightmost peak (19.3 MJ/mol) is the 1s subshell with 2 electrons, the next (1.36 MJ/mol) is 2s with 2 electrons, and the leftmost (0.80 MJ/mol) is 2p with 1 electron: $1s^2 2s^2 2p^1$, which is boron. A and B have only two subshells occupied, so they would show two peaks, not three. D ($1s^2 2s^2 2p^2$) would show the same three peak positions but the leftmost peak would be twice as tall relative to the others.

4. B. Across a period, protons are added to the nucleus while new electrons enter the same shell, so shielding barely changes and the effective nuclear charge rises. The stronger attraction pulls the electron cloud inward. A describes what happens down a group, not across a period. C reverses the ionization energy trend, which increases across a period. D misidentifies the mechanism: added protons increase attraction for the electrons, and electron-electron repulsion is not what shrinks the radius.

One-Page Recall Check

- Define the mole and state Avogadro's number with units.
- Explain the relationship between amu and g/mol using water as an example.
- Convert 36.0 g of H₂O to molecules, showing every unit cancellation.
- Describe what each peak in a mass spectrum represents and how to read relative abundance.
- Calculate the average atomic mass of chlorine from its two isotopes.
- State the law of definite proportions and explain why it does not apply to mixtures.
- Find the empirical formula of a compound that is 40.0% C, 6.7% H, and 53.3% O.
- Explain how a known molar mass turns an empirical formula into a molecular formula.
- Distinguish a pure substance from a mixture and give an example of each.
- Write the electron configuration of sodium and identify its valence electron.
- State Coulomb's law and use it to explain why the second ionization energy exceeds the first.
- Explain shielding and effective nuclear charge in your own words.
- Read a PES spectrum: what do peak position and peak height each tell you?
- Explain the atomic radius trend across a period and down a group using Coulomb's law.
- Explain why cations are smaller and anions larger than their parent atoms.
- Predict the ionic charge of magnesium and write the formula of magnesium chloride.

Study this unit in Rycal. Drill the 34 flashcards for Atomic Structure and Properties, then test yourself with AP-style questions on the mole, mass spectra, PES, and periodic trends at rycal.web.app/apchem. 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

Mole, Avogadro's number, Molar mass, Atomic mass unit (amu), Dimensional analysis, Formula unit, Mass spectrum, Isotope, Relative abundance, Average atomic mass, Law of definite proportions, Elemental analysis, Empirical formula, Molecular formula, Pure substance, Mixture, Elemental composition of a mixture, Proton, neutron, and electron, Electron configuration, Aufbau principle, Coulomb's law, Valence electrons, Ionization energy, Effective nuclear charge (shielding), Photoelectron spectroscopy (PES), Binding energy, Periodicity, Atomic radius trend, Ionic radius, Electron affinity, Electronegativity, Shell model, Predicting ionic charge, Analogous compounds.

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