

Unit 3: Intermolecular Forces and Properties

Unit 3 explains why matter behaves the way it does. It covers the forces between particles, how those forces set the properties of solids, liquids, and gases, the gas laws and kinetic molecular theory, solutions and how to separate them, and the spectroscopy chemists use to measure concentration.

AP Chemistry

Intermolecular Forces and Properties

About 14 minutes to read

How to use this guide

Read it in order the first time because the topics build on each other. Intermolecular forces explain boiling points and vapor pressure, which explain the solid types, which set up the gas laws, and the gas laws lead into kinetic molecular theory and its limits. The solutions and spectroscopy sections at the end apply the same ideas about particle interactions to mixtures and measurement.

After the first read, use the trap boxes and the confusion table to review the distinctions the exam tests most often. Finish with the practice questions, then work through the recall check on the last page out loud and note anything you cannot explain yet.

What this unit is worth. Unit 3 is about 18 to 22 percent of the AP Chemistry exam, the largest share of any unit. It also feeds into later units. Intermolecular forces return in solutions, acids and bases, and organic chemistry, and the gas laws underpin thermodynamics.

3.1 Intermolecular Forces

Intermolecular forces (IMFs) are the attractions between molecules, or between ions and molecules. They are much weaker than the covalent bonds holding a molecule together, but they control boiling point, melting point, vapor pressure, viscosity, and solubility. Almost every property in this unit traces back to which IMFs are present and how strong they are.

FORCE	WHAT IT IS
London dispersion forces (LDF)	Present between all molecules. They come from temporary, fluctuating dipoles in electron clouds. They strengthen with more electrons, larger contact area, and greater polarizability, and they are often the strongest IMF between large molecules.
Dipole-dipole forces	Attractions between the permanent dipoles of two polar molecules.
Hydrogen bonding	An especially strong dipole-dipole interaction. It happens when hydrogen bonded to N, O, or F is attracted to N, O, or F on another molecule, or on another part of the same molecule.
Ion-dipole forces	Attractions between an ion and a polar molecule, as when water molecules surround Na^+ and Cl^- as table salt dissolves.
Dipole-induced dipole forces	A polar molecule induces a temporary dipole in a neighboring nonpolar molecule, and the two attract. Strength depends on the size of the dipole and the polarizability of the nonpolar molecule.

Trap. Van der Waals forces is the general term for all intermolecular forces, not a synonym for London dispersion forces. LDFs are only one type. If a question asks for the van der Waals forces in a sample, it wants the full list, not just dispersion.

Trap. Hydrogen bonding needs H bonded to N, O, or F specifically. H_2S is polar, but sulfur is not N, O, or F, so H_2S has dipole-dipole forces and no hydrogen bonding. That is why water boils at 100°C while H_2S boils at -60°C .

Polarizability is how easily an electron cloud is distorted. It increases with more electrons and larger atoms, and pi bonding enhances it. This is why LDF strength tracks molar mass down a group. Iodine (I_2) is a solid at room temperature while fluorine (F_2) is a gas, because the larger electron cloud of I_2 is far more polarizable and the dispersion forces are far stronger.

In large biomolecules and polymers, **noncovalent interactions** happen between different molecules and between different regions of the same molecule. A protein folds into its working shape because of thousands of these weak attractions, so the molecule's shape, and therefore its function, is largely set by them.

3.2 Properties of Solids

The **boiling point** of a substance rises with the strength of its intermolecular forces, because more energy is needed to pull the particles apart. **Vapor pressure** is the pressure of the vapor in equilibrium with its liquid or solid, and it runs the other way. Stronger IMFs hold particles in the liquid, so fewer escape and the vapor pressure is lower. High boiling point and low vapor pressure go together.

SOLID TYPE	PROPERTIES
Ionic solids	Made of ions in a lattice. Low vapor pressure, high melting and boiling points, brittle, and they conduct electricity only when molten or dissolved in water, never as solids.
Covalent network solids	Nonmetal or metalloid atoms linked by covalent bonds in a continuous network. Three-dimensional networks like diamond, silicon dioxide, and silicon carbide are extremely hard with very high melting points. Graphite is a two-dimensional layered network, and it is soft because the layers slide past each other.
Molecular solids	Discrete molecules held together by weak IMFs. Low melting points, and they do not conduct electricity.
Metallic solids	Metal cations in a sea of mobile delocalized electrons. Good conductors of heat and electricity, malleable, and ductile. Alloys are less malleable than pure metals because the foreign atoms distort the lattice and block the layers from sliding.

Trap. Do not sort solids by hardness alone. Diamond and table salt are both hard, but diamond is a covalent network solid that melts above 3500°C while salt is an ionic solid that melts at 801°C and conducts when molten. Classify by particle type and bonding first, then predict the properties.

3.3 Solids, Liquids, and Gases

A **crystalline solid** has its particles arranged in a regular, repeating three-dimensional pattern, like the lattice of sodium chloride. An **amorphous solid** has no regular arrangement. Glass and many plastics are amorphous, which is why they soften over a range of temperatures instead of melting sharply.

The **molar volumes of solids and liquids** are typically similar for a given substance. In both phases the particles stay in close contact at all times, so condensing a liquid to a solid barely changes the volume per mole. The big volume change comes when a liquid becomes a gas.

3.4 Ideal Gas Law

The **ideal gas law**, $PV = nRT$, relates the pressure, volume, amount in moles, and temperature of an ideal gas. With pressure in atmospheres, volume in liters, and temperature in Kelvin, $R = 0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$. Temperature must be in Kelvin. A Celsius temperature plugged into $PV = nRT$ gives a wrong answer every time, and it is one of the most common arithmetic errors on the exam.

For a 2.00 mol sample at 300 K in a 10.0 L container: $P = nRT / V = (2.00)(0.08206)(300) / (10.0) = 4.92 \text{ atm}$. Notice the units cancel to leave atmospheres, which is a quick check that the setup is right.

Dalton's law of partial pressures says the total pressure of a gas mixture equals the sum of the **partial pressures** of the individual gases: $P_{\text{total}} = P_A + P_B + P_C + \dots$. Each partial pressure is the pressure that gas would exert on its own. The **mole fraction** of a component is its moles divided by the total moles, $X_A = n_A / n_{\text{total}}$, and the partial pressure follows from it: $P_A = X_A \times P_{\text{total}}$. If a 2.00 atm mixture is 25 percent oxygen by moles, the partial pressure of oxygen is $0.25 \times 2.00 \text{ atm} = 0.50 \text{ atm}$.

Trap. Partial pressure depends on mole fraction, not on the identity of the gas. At the same mole fraction, helium and carbon dioxide contribute the same partial pressure. Students sometimes weight by molar mass, but Dalton's law counts particles, not mass.

3.5 Kinetic Molecular Theory

Kinetic molecular theory (KMT) connects the macroscopic behavior of gases to particle motion. Its postulates are that gas particles move randomly in straight lines, their collisions are elastic, and their average kinetic energy is proportional to the Kelvin temperature. For one particle, $KE = \frac{1}{2}mv^2$, and the **average kinetic energy of gas particles** in a sample is set by temperature alone.

The **Maxwell-Boltzmann distribution** shows the spread of particle kinetic energies at a given temperature. As temperature rises, the peak shifts toward higher energy, the average energy rises, and the distribution spreads out and flattens. The area under the curve stays the same because the number of particles does not change.

Trap. At the same temperature, all gases have the same average kinetic energy, but not the same average speed. Since $KE = \frac{1}{2}mv^2$, lighter molecules must move faster to carry the same energy. Helium effuses faster than nitrogen at the same temperature for exactly this reason.

3.6 Deviation from Ideal Gas Law

Real gases follow $PV = nRT$ only approximately. **Deviations from ideal gas behavior** appear when intermolecular attractions matter, which happens at low temperatures near the conditions where the gas would condense, or when the volume of the particles themselves is significant, which happens at high pressure. The ideal gas law assumes point particles with no attractions, so both assumptions fail at the extremes.

3.7 Solutions and Mixtures

A **homogeneous mixture**, also called a solution, has uniform properties throughout the sample.

Solutions are not limited to liquids. Air is a gaseous solution and brass is a solid solution. A

heterogeneous mixture is one whose properties vary depending on where in the sample you look, like sand stirred into water.

Molarity is concentration in moles of solute per liter of solution: $M = n_{\text{solute}} / L_{\text{solution}}$. The denominator is liters of solution, not liters of solvent. To make 2.00 L of 0.500 M NaCl, you need 1.00 mol of NaCl, and you add water until the total solution volume reaches 2.00 L.

Trap. Molarity uses the volume of the finished solution. Dissolving 1.00 mol of solute in 1.00 L of water does not give a 1.00 M solution, because the final volume exceeds 1.00 L once the solute is added. Volumetric glassware exists to fix exactly this mistake.

3.8 Representations of Solutions

Particulate representations of solutions are particle-level diagrams. They show the relative concentrations of the components and the interactions between solute and solvent particles. On the exam, read them literally. Count the particles to compare concentrations, and look at how the solvent particles orient around the solute to identify ion-dipole or hydrogen-bonding interactions.

3.9 Separation of Solutions and Mixtures

Filtration passes a mixture through a porous barrier. It separates solids from liquids or gases, but it cannot separate a dissolved solute from a liquid solution. The dissolved ions or molecules are far smaller than the pores.

Chromatography, in paper, thin-layer, or column form, separates components by how strongly they interact with a mobile phase versus a stationary phase. The resulting chromatogram lets you infer relative polarities. A component that travels farther up the paper interacted more strongly with the mobile phase and less strongly with the stationary phase.

Distillation exploits differences in vapor pressure, which is the same as differences in boiling point, between the components of a mixture. The more volatile component vaporizes first, condenses, and is collected separately.

Trap. Match the method to the property difference. Filtration needs a particle-size difference, distillation needs a boiling-point difference, and chromatography needs a difference in attraction to the two phases. Filtering salt water leaves you with salt water.

3.10 Solubility

Like dissolves like is the rule for solubility. A substance tends to dissolve in a solvent with similar intermolecular forces: polar solutes in polar solvents, nonpolar solutes in nonpolar solvents. Ethanol mixes with water because both can hydrogen bond. Oil does not, because breaking water's hydrogen bonds to accommodate nonpolar oil molecules costs more energy than the weak dispersion interactions return.

3.11–3.12 Spectroscopy and the Electromagnetic Spectrum

Different regions of the **electromagnetic spectrum** drive different molecular changes. Microwaves cause rotational transitions, infrared causes vibrational transitions, and ultraviolet and visible light cause electronic transitions. Higher-frequency light carries more energy per photon, so electronic transitions need UV or visible light while rotations need only microwaves.

When a species absorbs or emits a photon, its energy changes by exactly the photon's energy in a **photon absorption and emission** event. Two equations govern the numbers. The **wave equation**, $c = \lambda\nu$, relates the speed of light to wavelength and frequency. **Planck's equation**, $E = h\nu$, gives the energy of a photon from its frequency, with $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$. Combine them to get energy straight from wavelength: $E = hc / \lambda$.

Trap. Frequency and wavelength are inversely related through $c = \lambda\nu$. Doubling the frequency halves the wavelength and doubles the photon energy. Students sometimes treat them as independent, but fixing one fixes the other two.

3.13 Beer-Lambert Law

The **Beer-Lambert law**, $A = \epsilon bc$, says the absorbance of a solution is proportional to three things multiplied together: the **molar absorptivity (ϵ)**, the path length b , and the concentration c . Molar absorptivity describes how intensely a chemical species absorbs light of a specific wavelength. Because absorbance is directly proportional to concentration, a chemist can measure absorbance and read concentration off a calibration curve.

The proportionality makes the math simple. If a solution with concentration 0.100 M has an absorbance of 0.45, doubling the concentration to 0.200 M doubles the absorbance to 0.90, as long as the path length and wavelength stay the same. Halving the concentration halves it.

A **spectrophotometer** is the instrument that measures absorbance. It is set to the wavelength of maximum absorbance for the species being measured, because that is where a small change in concentration produces the largest change in signal, giving the most sensitive measurements.

Trap. The Beer-Lambert law is linear only within limits. At very high concentrations the relationship curves because solute particles start interacting with each other, and stray light in the instrument flattens the readings. Calibration curves are built from standards in the linear range for exactly this reason.

Confusions That Cost Points

PAIR	HOW TO KEEP THEM STRAIGHT
LDF vs dipole-dipole vs hydrogen bonding	LDFs are in everything and grow with electron count. Dipole-dipole needs permanent dipoles. Hydrogen bonding needs H bonded to N, O, or F. Check for N, O, F before claiming hydrogen bonding.
Van der Waals vs LDF	Van der Waals is the umbrella term for all IMFs. LDF is one type under it. Never use them interchangeably.
Boiling point vs vapor pressure	Stronger IMFs raise the boiling point and lower the vapor pressure. They move in opposite directions.
Ionic vs molecular solids	Ionic solids conduct when molten or dissolved and have high melting points. Molecular solids have low melting points and never conduct. The particle type decides.
Diamond vs graphite	Both are covalent network solids, but diamond is a 3D network and extremely hard, while graphite is 2D layers that slide, so it is soft.
Ideal vs real gas conditions	Ideal behavior holds at high temperature and low pressure. Real gases deviate at low temperature (IMFs matter) and high pressure (particle volume matters).
Molarity denominator	Liters of solution, not liters of solvent. The solute contributes volume.
Filtration vs distillation vs chromatography	Filtration separates by particle size, distillation by boiling point, chromatography by attraction to the mobile versus stationary phase.
Absorbance vs concentration	Directly proportional through $A = \epsilon bc$. Double the concentration and the absorbance doubles, within the linear range.

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. Which of the following substances has the highest boiling point?

- A. CH_4
- B. H_2S
- C. H_2O
- D. He

2. A 2.00 mol sample of an ideal gas at 300 K occupies a 10.0 L container. The pressure of the gas is closest to

- A. 0.49 atm
- B. 4.9 atm
- C. 49 atm
- D. 0.20 atm

3. A real gas deviates most from ideal behavior under which of the following conditions?

- A. High temperature and low pressure
- B. Low temperature and high pressure
- C. High temperature and high pressure
- D. Low temperature and low pressure

4. A solution has an absorbance of 0.45 at a particular wavelength. The solution is diluted to half its original concentration, with the same path length and wavelength. The new absorbance is closest to

- A. 0.90
- B. 0.225
- C. 0.45
- D. 0.11

Answer Key

1. C. H_2O hydrogen bonds, since its hydrogen atoms are bonded to oxygen. That is the strongest IMF in the list. H_2S is polar but sulfur is not N, O, or F, so it has only dipole-dipole forces plus LDFs. CH_4 and He have only LDFs, and He, with two electrons, has the weakest. A and D ignore hydrogen bonding entirely. B mistakes any polar molecule for a hydrogen-bonding one.

2. B. $P = nRT / V = (2.00)(0.08206)(300) / (10.0) = 4.92 \text{ atm}$, closest to 4.9 atm. A comes from using 30 K or misplacing a decimal. C comes from forgetting to divide by the volume. D may come from inverting the expression (V / nRT).

3. B. Low temperature lets intermolecular attractions matter because the particles move slowly near condensation conditions, and high pressure makes the particles' own volume significant. A describes the conditions where gases behave most ideally. C gets pressure right but temperature wrong. D gets temperature right but pressure wrong; both extremes are needed.

4. B. The Beer-Lambert law, $A = \epsilon bc$, makes absorbance directly proportional to concentration. Halving the concentration halves the absorbance: $0.45 / 2 = 0.225$. A doubles instead of halving. C treats absorbance as independent of concentration. D quarters it, which would require halving twice.

One-Page Recall Check

- List the five types of intermolecular forces and state when each one applies.
- Explain why I_2 is a solid at room temperature while F_2 is a gas.
- State the N, O, F rule for hydrogen bonding and give one example that fails it.
- Explain the difference between van der Waals forces and London dispersion forces.
- Predict the relative boiling points and vapor pressures of two substances from their IMFs.
- Classify a solid as ionic, covalent network, molecular, or metallic from its properties.
- Explain why graphite is soft even though diamond is extremely hard.
- Distinguish crystalline from amorphous solids.
- Use $PV = nRT$ with correct units, including Kelvin temperature.
- Apply Dalton's law with mole fractions to find a partial pressure.
- State the postulates of kinetic molecular theory.
- Describe how the Maxwell-Boltzmann distribution changes with temperature.
- Explain why lighter gases move faster than heavier gases at the same temperature.
- State the two conditions under which real gases deviate from ideal behavior.
- Distinguish homogeneous from heterogeneous mixtures and give a non-liquid example of each.
- Calculate molarity and explain why the denominator is liters of solution.
- Choose filtration, chromatography, or distillation for a given separation and justify the choice.
- State the like-dissolves-like rule in terms of intermolecular forces.
- Match microwave, infrared, and UV/visible light to the molecular transitions each causes.
- Use $c = \lambda\nu$ and $E = h\nu$ together to find a photon's energy from its wavelength.
- Use $A = \epsilon bc$ to relate absorbance and concentration, and explain the spectrophotometer setting.

Study this unit in Rycal. Drill the key terms from this unit as flashcards, then test yourself with AP-style practice questions at rycal.web.app/apchem. If you have a test date, add it in the Test Planner so your review schedule builds around it.

Key terms for this unit

London dispersion forces (LDF), Polarizability, Dipole-induced dipole forces, Dipole-dipole forces, Ion-dipole forces, Hydrogen bonding, van der Waals forces, Noncovalent interactions in biomolecules, Boiling point and IMFs, Vapor pressure, Ionic solids, Covalent network solids, Molecular solids, Metallic solids (properties), Crystalline solid, Amorphous solid, Molar volumes of solids and liquids, Ideal gas law, Partial pressure, Dalton's law of partial pressures, Mole fraction, Kinetic molecular theory (KMT), Maxwell-Boltzmann distribution, Average kinetic energy of gas particles, Deviations from ideal gas behavior, Homogeneous mixture (solution), Heterogeneous mixture, Molarity, Particulate representations of solutions, Filtration, Chromatography, Distillation, Like dissolves like, Electromagnetic spectrum and transitions, Photon absorption and emission, Wave equation, Planck's equation, Beer-Lambert law, Spectrophotometer, Molar absorptivity (ϵ).

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