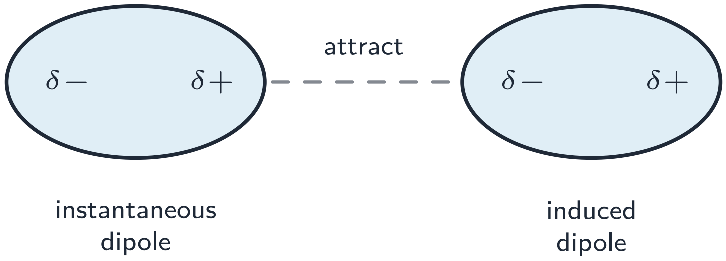
# Properties of Substances and Mixtures

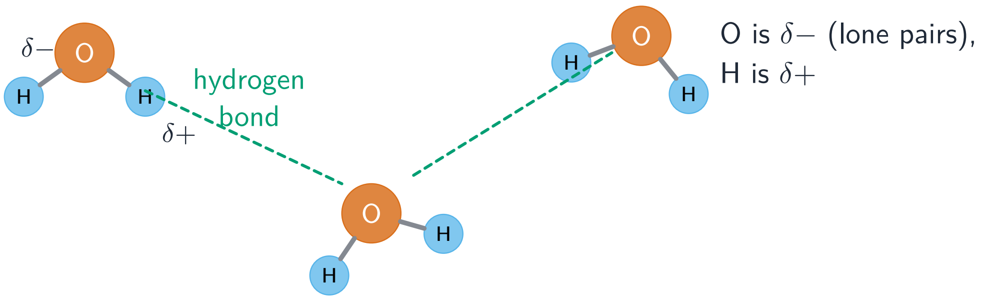
AP Chemistry

## Intermolecular and Interparticle Forces

**Intermolecular forces** 分子间作用力 (IMFs) are attractions **between** molecules –much weaker than bonds, but they set melting/boiling points. From weakest to strongest:



*London dispersion: an instantaneous dipole induces a dipole in a neighbour*



*Hydrogen bonding: an H on N/O/F is attracted to a lone pair on another molecule*

- **London dispersion forces** 伦敦色散力: present in all molecules; stronger for larger, more polarizable electron clouds.
- **Dipole–dipole** 偶极-偶极: between polar molecules.
- **Hydrogen bonding** 氢键: a strong dipole force when H is bonded to N, O, or F.

Stronger IMFs mean higher boiling points and lower vapor pressure. This is why water (18 g/mol, hydrogen-bonded) boils at 100 °C while methane (16 g/mol, dispersion only) boils at $-162\,°\text{C}$.

## Properties of Solids

A solid’s properties reflect the particles and forces holding it: **ionic**, **covalent-network** (like diamond, very hard, high-melting), **metallic**, and **molecular** solids (held by weak

IMFs, soft, low-melting). Matching a solid's properties to its structure is a common exam task.

**giant ionic**
NaCl: high m.p.,
conducts when molten

**simple molecular**
$I_2$, ice: low m.p.,
weak forces between

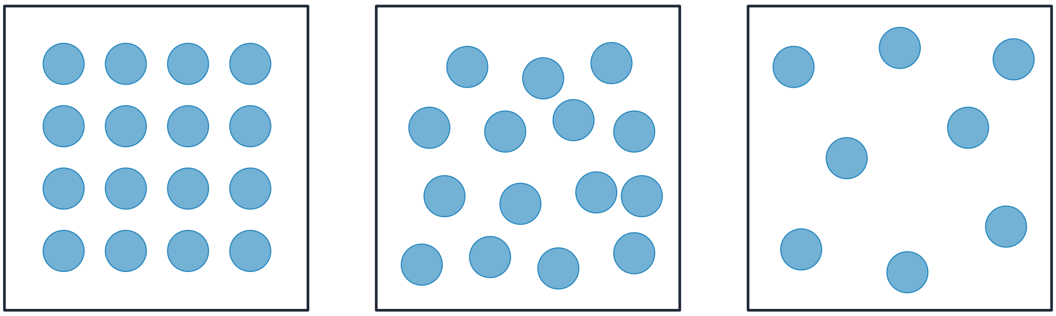
**giant covalent**
diamond, $SiO_2$: very
high m.p., insulating

**giant metallic**
copper: high m.p.,
conducts (solid + molten)

*The four solid structures: giant ionic, simple molecular, giant covalent, and metallic*

## Solids, Liquids, and Gases

The three states differ in how tightly particles are held. Rising **temperature** raises average kinetic energy; when it overcomes the attractions, the substance melts or boils. Gases are mostly empty space, so they are compressible and fill their container.



**solid**
regular, touching
vibrate in place

**liquid**
close, random
slide past each other

**gas**
far apart, random
move fast

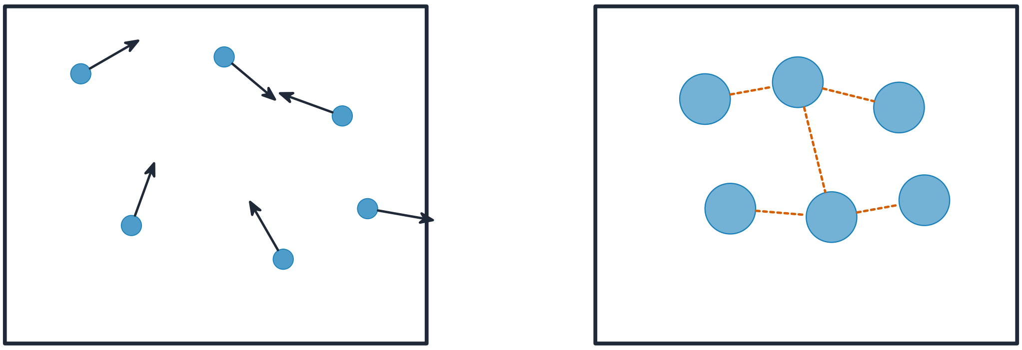
*Particles are packed in a solid, close but mobile in a liquid, and far apart in a gas*

# The Ideal Gas Law

An **ideal gas** obeys

$$PV = nRT,$$

linking pressure, volume, moles, and absolute temperature. Use it to find any one quantity from the others, or (holding some constant) to predict how a gas responds to a change.



**ideal gas** (model)
point particles, *no* forces

**real gas**: high $p$, low $T$
real size + attractions (dashed)

*An ideal gas is a model of point particles with no forces between them*

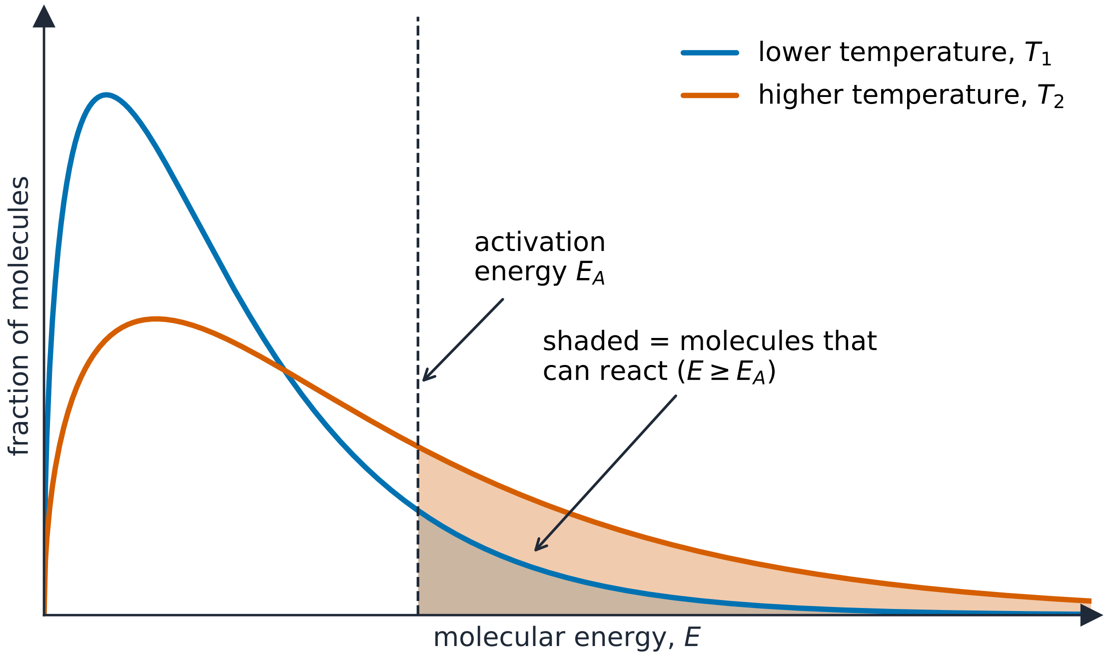
**Worked example.** How many moles of gas fill a 2.0 L container at 300 K and 1.5 atm? Using $R = 0.0821$ L atm/(mol K),

$$n = \frac{PV}{RT} = \frac{1.5 \times 2.0}{0.0821 \times 300} = 0.12 \text{ mol}.$$

Always use **kelvin** for $T$ and match the units of $R$ to your pressure and volume.

# Kinetic Molecular Theory

**Kinetic molecular theory** 分子运动论 explains gas behavior: particles are tiny, in constant random motion, with negligible volume and no attractions, and collisions are elastic. **Temperature** is proportional to average kinetic energy, so at a given temperature lighter molecules move faster (Graham's law of effusion).



*The Maxwell-Boltzmann distribution of molecular speeds shifts right when heated*

## Deviation from the Ideal Gas Law

**Real gases** deviate from ideal behavior at **high pressure** and **low temperature**, where molecules are close enough that their real volume and their attractions matter. Attractions lower the pressure below ideal; molecular volume raises it.

## Solutions and Mixtures

A **solution** 溶液 is a homogeneous mixture of a **solute** 溶质 dissolved in a **solvent** 溶剂. Concentration is usually **molarity** 摩尔浓度:

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}.$$

**Dilution** conserves moles: $M_1V_1 = M_2V_2$.

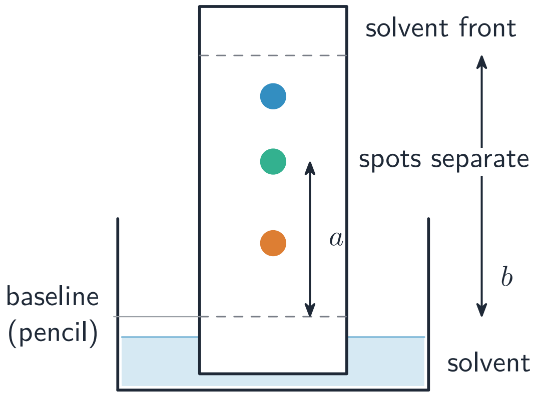
**Worked example.** What volume of water must you add to 50 mL of 6.0 M HCl to make it 2.0 M? The moles of HCl are unchanged, so $M_1V_1 = M_2V_2$ gives the final volume $V_2 = \dfrac{M_1V_1}{M_2} = \dfrac{6.0 \times 50}{2.0} = 150$ mL. You therefore **add** $150 - 50 = 100$ mL of water.

## Representations of Solutions

A **particulate diagram** shows the solute and solvent particles. For an ionic solute, show it fully **dissociated** into separate ions surrounded by solvent; count particles to reason about concentration and conductivity.

## Separation of Solutions and Mixtures

Because a mixture’s components keep their properties, physical methods separate them: **filtration** (by particle size), **distillation** (by boiling point), and **chromatography** 色谱法 (by how strongly each component sticks to a stationary phase versus moving with a solvent).



$$R_f = \frac{a}{b} = \frac{\text{spot}}{\text{solvent}}$$

*Paper chromatography separates a mixture as the solvent rises up the paper*

## Solubility

**Solubility** 溶解度 is how much solute dissolves. ”Like dissolves like”: polar (and ionic) solutes dissolve in polar solvents; nonpolar in nonpolar. Dissolving happens when solute–solvent attractions are comparable to the attractions being broken.

## Spectroscopy and the Electromagnetic Spectrum

**Spectroscopy** 光谱学 studies how matter absorbs or emits light. Different regions of the **electromagnetic spectrum** probe different changes: microwaves (rotation), infrared (bond vibrations), ultraviolet–visible (electron transitions). The light absorbed reveals structure.

## Properties of Photons

Light is carried by **photons** 光子, each with energy $E = h\nu = \dfrac{hc}{\lambda}$. Higher frequency (shorter wavelength) means higher energy. A molecule absorbs a photon only when its energy matches an allowed energy gap.

# The Beer-Lambert Law

The **Beer–Lambert law** 比尔-朗伯定律 relates how much light a solution absorbs to its concentration:

$$A = \varepsilon\, b\, c,$$

where $A$ is absorbance, $\varepsilon$ the molar absorptivity, $b$ the path length, and $c$ the concentration. Since $A$ is proportional to $c$, measuring absorbance is a fast way to find an unknown concentration.

**Worked example.** A dye has molar absorptivity $\varepsilon = 2000$ L/(mol cm); in a 1.0 cm cell a sample reads absorbance $A = 0.40$. Its concentration is $c = \dfrac{A}{\varepsilon b} = \dfrac{0.40}{2000 \times 1.0} = 2.0 \times 10^{-4}$ M. Because $A \propto c$, a solution twice as concentrated would read $A = 0.80$ – the basis of a calibration curve.

# Exam tips

- **Intermolecular forces** (dispersion < dipole–dipole < hydrogen bonding) set boiling points —they are much weaker than the bonds inside a molecule.
- Boiling breaks the forces **between** molecules, not the covalent bonds within them.
- Use $PV = nRT$ with temperature in **kelvin** and $R$ matched to your pressure/volume units.
- For solutions use molarity $M = \text{mol/L}$; dilution conserves moles, so $M_1V_1 = M_2V_2$.
- ”Like dissolves like” —polar/ionic solutes dissolve in polar solvents, non-polar in non-polar.