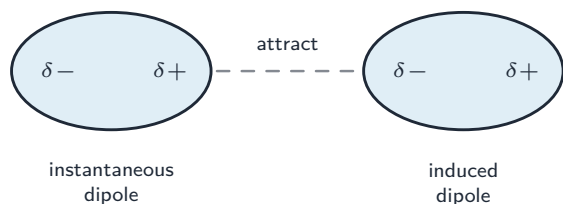


Properties of Substances and Mixtures

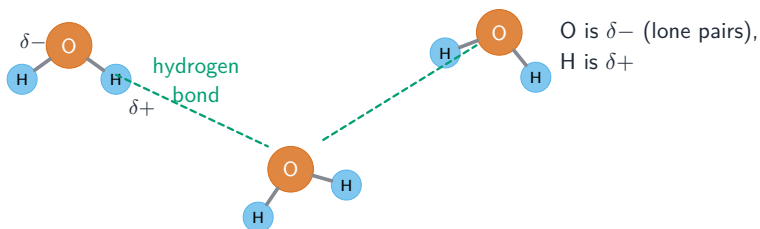
AP Chemistry

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; they arise from Coulombic attraction between **temporary, fluctuating dipoles**, and are stronger for larger, more polarizable electron clouds - often the strongest net force between large molecules.
- **Dipole-dipole** 偶极-偶极: between polar molecules. Their strength depends on the size of the dipoles **and their relative orientation** - a $\delta+$ end lining up with a neighbour's $\delta-$ end attracts, so these act *in addition* to dispersion and make polar molecules stickier than nonpolar ones of similar size.
- **Hydrogen bonding** 氢键: a strong dipole force when H is bonded to N, O, or F.

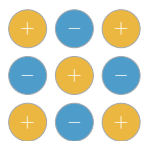
- **Ion-dipole** 离子-偶极: between an **ion** and a polar molecule (the ion pulls on the oppositely-charged end of the dipole). These are the **strongest** of the four here (stronger even than hydrogen bonds), and are exactly what lets water dissolve an ionic solid - each Na^+ is surrounded by the $\delta-$ oxygen ends of water molecules.

The whole ladder is understood **qualitatively** by looking at the **sign of the partial charges**: a larger or better-aligned partial charge, or a full ionic charge, gives a stronger attraction. 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°C .

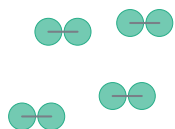
Properties of Solids

A solid's properties reflect the particles and forces holding it: **ionic**, **covalent-network** (a 3D network like diamond, very hard and high-melting; **graphite** is a layered exception — high-melting but soft, because its 2D layers slide over one another), **metallic**, and **molecular** solids (held by weak IMFs, soft, low-melting). Matching a solid's properties to its structure is a common exam task.

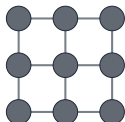
Metallic solids conduct electricity and heat and are **malleable** 可锻的 and **ductile** 可延展的, all because their **free valence electrons** 自由价电子 move easily and let the metal ions slide past one another without breaking the bonding. Solids are also either **crystalline** 晶体 (particles in a regular, repeating 3-D arrangement) or **amorphous** 非晶体 (no long-range order, like glass).



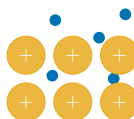
giant ionic
NaCl: high m.p.,
conducts when molten



simple molecular
I₂, ice: low m.p.,
weak forces between



giant covalent
diamond, SiO₂: very
high m.p., insulating

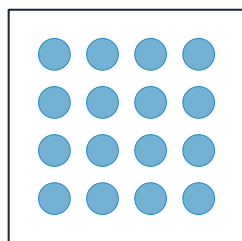


giant metallic
copper: high m.p.,
conducts (solid + molten)

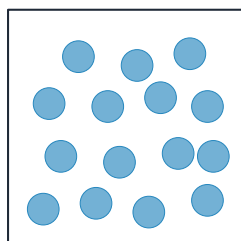
The four solid structures: giant ionic, simple molecular, giant covalent, and metallic

Solids, Liquids, and Gases

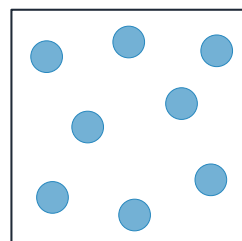
The three states differ in how tightly particles are held. Rising **temperature** raises average kinetic energy; when it overcomes the attractions, the substance melts or boils. Gases are mostly empty space, so they are compressible and fill their container.



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



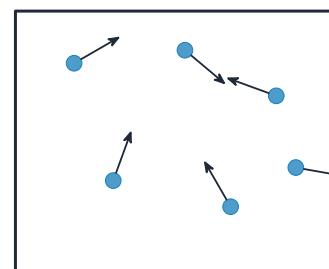
Steam rising from boiling water: gases expand to fill space and obey the ideal gas law at high T and low P

The Ideal Gas Law

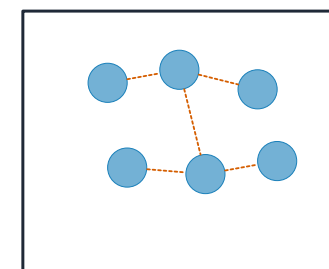
An **ideal gas** obeys

$$PV = nRT,$$

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

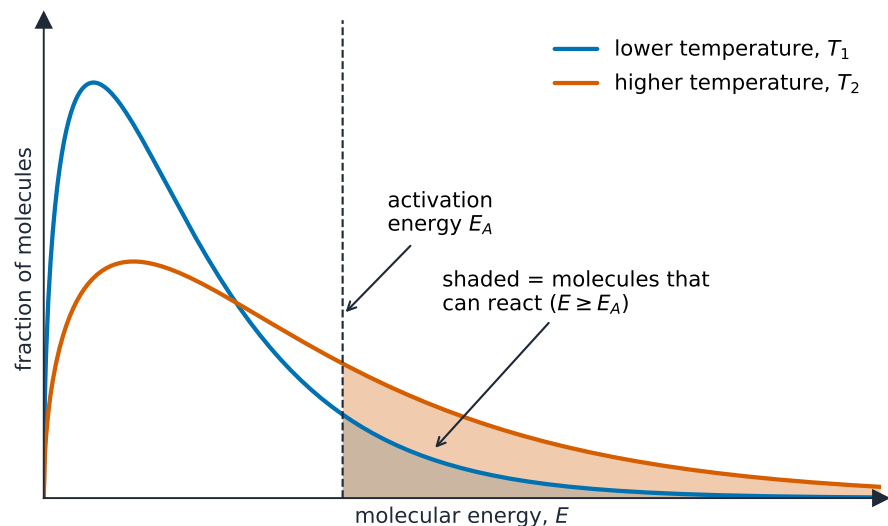
1.5 atm? Using $R = 0.0821 \text{ L atm}/(\text{mol K})$,

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

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: $M_1V_1 = M_2V_2$.

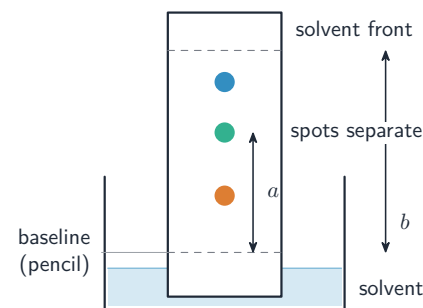
Worked example. What volume of water must you add to 50 mL of 6.0 M HCl to make it 2.0 M? The moles of HCl are unchanged, so $M_1V_1 = M_2V_2$ gives the final volume $V_2 = \frac{M_1V_1}{M_2} = \frac{6.0 \times 50}{2.0} = 150 \text{ mL}$. You therefore **add** $150 - 50 = 100 \text{ mL}$ of water.

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.



Blue copper(II) sulfate crystals grown from solution: a saturated solution left to evaporate deposits its dissolved solid back out as regular crystals

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 $E = h\nu = \frac{hc}{\lambda}$. 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, ε 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 $\varepsilon = 2000 \text{ L}/(\text{mol cm})$; in a 1.0 cm cell a sample reads absorbance $A = 0.40$. Its concentration is $c = \frac{A}{\varepsilon b} = \frac{0.40}{2000 \times 1.0} = 2.0 \times 10^{-4} \text{ M}$. 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.