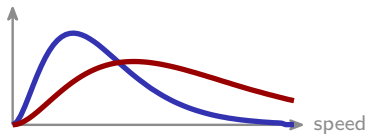


Properties of Substances and Mixtures

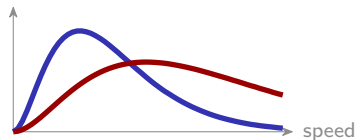
AP Chemistry ■ Unit 3

AP Chemistry



What we will cover

- 1 Intermolecular forces
- 2 Types of solids
- 3 States of matter and the gas law
- 4 Kinetic molecular theory
- 5 Solutions and solubility
- 6 Separation methods
- 7 Spectroscopy and Beer-Lambert



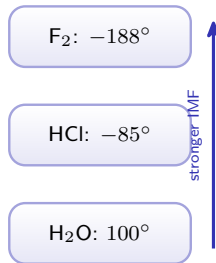


Intermolecular forces

IMFs are weak attractions *between* molecules – they set boiling points:

- **London dispersion** – all molecules; grows with size
- **dipole-dipole** – polar molecules
- **hydrogen bonding** – H bonded to N, O, F

H-bond > dipole > dispersion – but compare the **total**.



The ladder of intermolecular forces

London dispersion

present in **all** molecules – Coulombic attraction between **temporary, fluctuating dipoles**. Stronger with more electrons and a larger, more polarisable cloud

dipole–dipole –

between polar molecules; strength depends on the size of the dipoles **and their relative orientation**

hydrogen bonding

the strongest of these – H bonded to N, O or F

ion–dipole

an ion beside a polar molecule – what dissolves a salt in water

how to reason

the whole ladder is judged **qualitatively** from the **sign and size of the partial charges**, and how well they can align

Bigger or better-aligned partial charges, or a more polarisable electron cloud, means a stronger force – and a higher boiling point.

The intermolecular forces, weakest first

London dispersion forces

London dispersion forces act between *all* molecules: a momentary dipole induces one next door. Stronger with more electrons.

Dipole–dipole

Dipole–dipole forces act between permanently polar molecules, adding to the dispersion forces already there.

Hydrogen bonding

The special strong case: H bonded to N, O or F, attracted to a lone pair on another such atom.

Ion–dipole

Ion–dipole is stronger still, and it is what dissolving an ionic solid in water actually is.

Compare boiling points by asking which forces are present *and* how many electrons there are. Both matter.



Types of solids

A solid's properties follow from its particles and the force holding them:

Ionic : high melt, brittle, conducts molten

Covalent network : very high melt, hard (diamond)

Metallic : conducts, malleable

Molecular : low melt, soft (ice)

Network solids melt highest – you must break **covalent bonds**, not just weak IMFs.

Ordered or not, and one measurement

Crystalline

A **crystalline** solid has a repeating lattice, so it melts at one sharp temperature and cleaves along planes.

Amorphous

An **amorphous** solid – glass, many polymers – has no long-range order, so it softens over a range instead of melting.

Beer–Lambert law

The **Beer–Lambert law** $A = \epsilon bc$ makes concentration measurable: absorbance is proportional to path length and concentration.

Why it is used

A calibration curve of known concentrations turns one absorbance reading into an unknown's concentration.

The AP lab uses Beer–Lambert constantly. Plot A against c , and the unknown is read off the straight line.

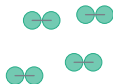
Types of solids, in detail

**giant ionic**

NaCl: high m.p.,
conducts when molten

**giant covalent**

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

**simple molecular**

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

**giant metallic**

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

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

Why each kind of solid behaves as it does

metallic

conducts electricity and heat, **malleable** and **ductile** – all because the **free valence electrons** are delocalised across the lattice

ionic

hard and brittle, high melting point; conducts only when molten or dissolved, because the ions must be free to move

covalent network

very high melting point – breaking it means breaking covalent bonds

molecular

low melting point – only the intermolecular forces are broken, not the bonds

the states

differ in how tightly particles are held; rising **temperature** raises average kinetic energy until it overcomes the attractions

Every property follows from **what has to be broken** and **what can move**. Answer those two questions and the property follows.



States of matter and the gas law

A balance of **kinetic energy** vs IMFs sets the state:

Ideal gas law

$$PV = nRT \quad (T \text{ in kelvin})$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Dalton : $P_{\text{tot}} = \sum P_i$, with $P_i = x_i P_{\text{tot}}$.



Gases fill volume; solids hold shape

Worked example – the ideal gas law

How many moles of gas fill a 2.0 L container at 300 K and 1.5 atm?

- $PV = nRT$, with
 $R = 0.0821 \text{ L atm}/(\text{mol K})$
- $n = \frac{PV}{RT} = \frac{(1.5)(2.0)}{(0.0821)(300)}$
- $= 0.12 \text{ mol}$

Check the units of R against the units you were given: with pressure in kPa you need a different R , and that mismatch is the commonest error in this calculation.



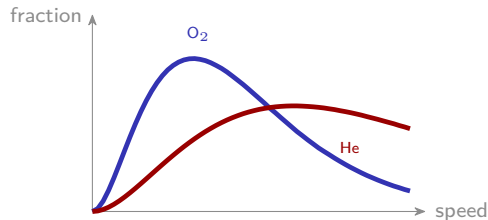
Kinetic molecular theory

KMT : gas particles are tiny points, in constant random motion, with elastic collisions and no attractions.

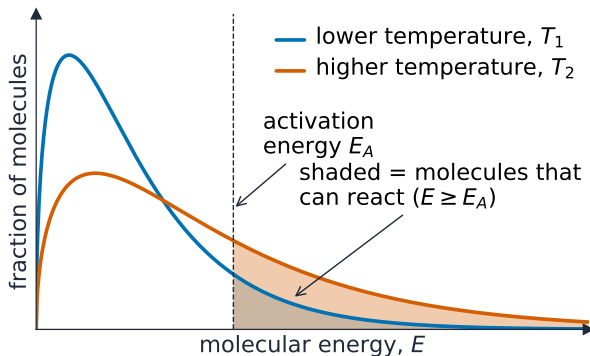
Temperature is average KE

$$\overline{KE} = \frac{3}{2}k_B T = \frac{1}{2}mv^2$$

Same T : lighter particles move **faster**.
Real gases stray at **high P , low T** .



The Maxwell-Boltzmann distribution of molecular



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



Solutions and solubility

A **solution** : **solute** dissolved in **solvent** .

Molarity and dilution

$$M = \frac{n}{V}, \quad M_1 V_1 = M_2 V_2$$

“Like dissolves like”: polar in polar,
nonpolar in nonpolar.



Solute and solvent make a solution

Drawing a particulate diagram of a solution

what it shows

the **solute** and **solvent** particles, at the particle level

an ionic solute

show it fully **dissociated** into **separate ions**, each surrounded by solvent molecules

the orientation

water's δ^- oxygen faces the cations, its δ^+ hydrogens face the anions

a molecular solute

show whole molecules, not fragments – sugar does not dissociate

a strong acid

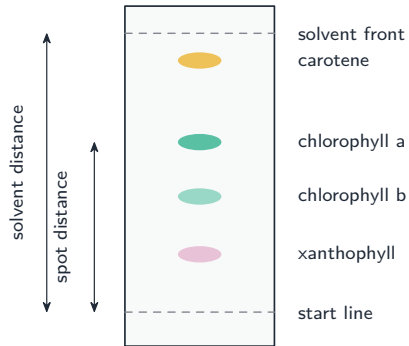
fully ionised: no intact HCl molecules should appear

the marks

come from correct **ratio, dissociation and orientation**, not from artistic quality

Count the particles you draw. A diagram showing the wrong stoichiometric ratio loses the mark however neat it is.

Separation of Solutions and Mixtures



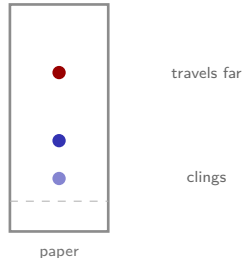
Paper chromatography separates a mixture as the solvent rises up the paper



Separation methods

Mixtures separate by **physical** differences:

- **chromatography** – different pull between a stationary and mobile phase
- **distillation** – different boiling points
- **filtration** – different particle sizes





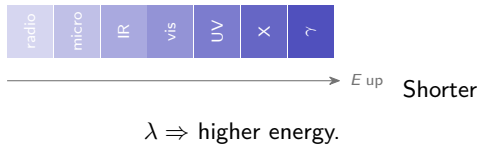
Spectroscopy and the EM spectrum

Absorbed light matches an energy gap:

Photon energy

$$E = h\nu = \frac{hc}{\lambda}$$

- microwave – **rotational**
- **infrared** – **vibrational**
- UV/visible – **electronic**



Properties of Photons

Light is carried by **photons** , each with energy $E = h\nu = \frac{hc}{\lambda}$.

Higher frequency (shorter wavelength) means higher energy.

The Beer-Lambert law

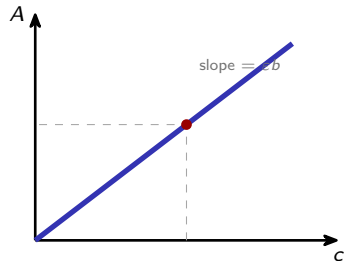
A coloured solution's **absorbance** is proportional to its concentration:

Beer-Lambert

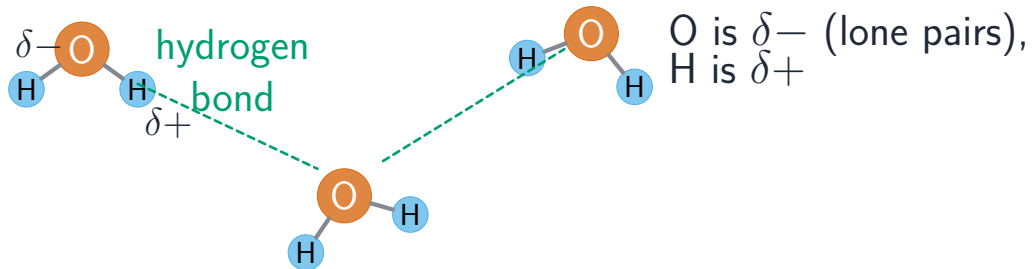
$$A = \epsilon bc$$

ϵ = molar absorptivity , b = path length .

A **calibration curve** of A vs c is a straight line – read an unknown off it.

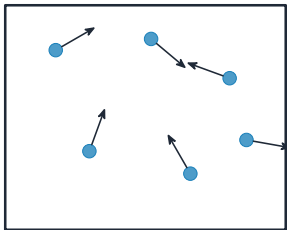


Intermolecular and Interparticle Forces

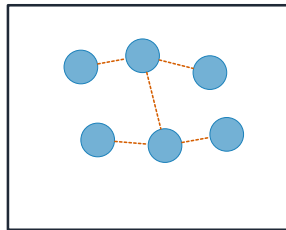


Hydrogen bonding: an H on N/O/F is attracted to a lone pair on another molecule

The Ideal Gas Law



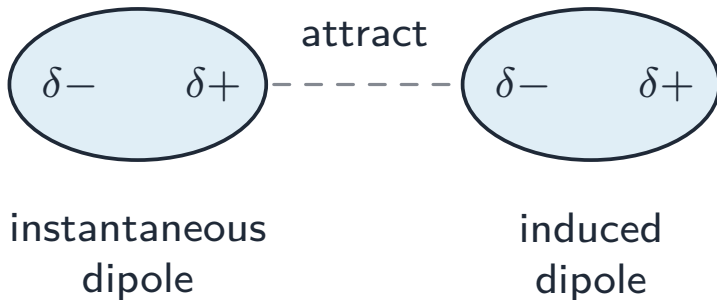
ideal gas (model)
point particles, *no* forces



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

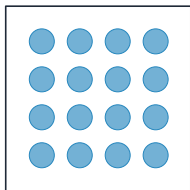
An ideal gas is a model of point particles with no forces between them

Intermolecular and Interparticle Forces, continued

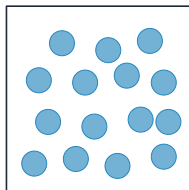


London dispersion: an instantaneous dipole induces a dipole in a neighbour

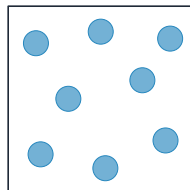
Solids, Liquids, and Gases

**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

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 $M_1 V_1 = M_2 V_2$.
- “Like dissolves like” – polar/ionic solutes dissolve in polar solvents, non-polar in non-polar.



Unit 3 – key takeaways

1

IMFs

H-bond > dipole > dispersion;
they set boiling points

2

Gases

$PV = nRT$; KMT – lighter particles move faster

3

Solutions

$M = n/V$; dilution $M_1 V_1 = M_2 V_2$; like dissolves like

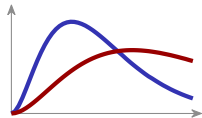
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Light

$E = hc/\lambda$; Beer-Lambert $A = \epsilon bc$

Check yourself

- 1 Which IMF gives water its unusually high boiling point?
- 2 At the same temperature, does He or O₂ move faster?
- 3 Dilute a solution – do the moles of solute change?
- 4 In Beer-Lambert, absorbance is proportional to what?



Answers 1. hydrogen bonding 2. He (lighter) 3. no (only concentration) 4. concentration