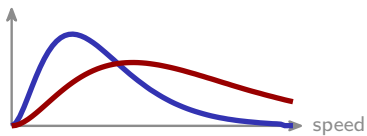
# Properties of Substances and Mixtures

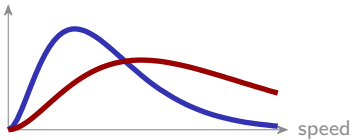
## AP Chemistry ▪ Unit 3

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



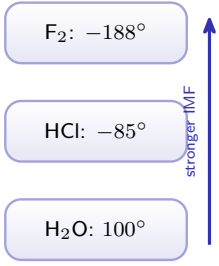
# 1 IMFs

# Intermolecular forces

**IMFs** 分子间作用力 are weak attractions *between* molecules – they set boiling points:

- **London dispersion** 伦敦色散力 – all molecules; grows with size
- **dipole-dipole** 偶极-偶极 – polar molecules
- **hydrogen bonding** 氢键 – H bonded to N, O, F

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



# The ladder of intermolecular forces

| | |
|---|---|
| **London dispersion 伦敦色散力** | present in all molecules – Coulombic attraction between temporary, fluctuating dipoles. Stronger with more electrons and a larger, more polarisable cloud |
| **dipole–dipole 偶极-偶极** | between polar molecules; strength depends on the size of the dipoles and their relative orientation |
| **hydrogen bonding 氢键** | 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.

# 2 Solids

# Types of solids

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

**Ionic** 离子型: high melt, brittle, conducts molten

**Covalent network** 共价网状: very high melt, hard (diamond)

**Metallic** 金属型: conducts, malleable

**Molecular** 分子型: low melt, soft (ice)

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

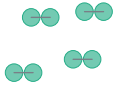
# Ordered or not, and one measurement

| | |
|---|---|
| **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** $A = \varepsilon bc$ 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



**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

# Why each kind of solid behaves as it does

| | |
|---|---|
| **metallic** | conducts electricity and heat, **malleable** 可锻的 and **ductile** 可延展的 – 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.

# 3 Gas law

# States of matter and the gas law

A balance of **kinetic energy** 动能 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** 道尔顿分压定律: $P_{tot} = \sum P_i$, with $P_i = x_i P_{tot}$.

*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 $R = 0.0821$ L atm/(mol K)
- $n = \dfrac{PV}{RT} = \dfrac{(1.5)(2.0)}{(0.0821)(300)}$
- $= 0.12$ 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.

4

# KMT

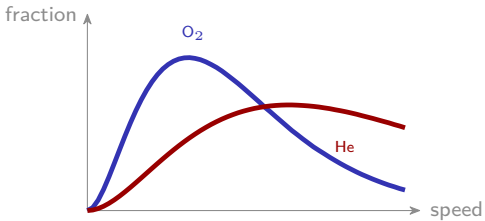
# Kinetic molecular theory

**KMT** 分子运动论: gas particles are tiny points, in constant random motion, with elastic collisions and no attractions.

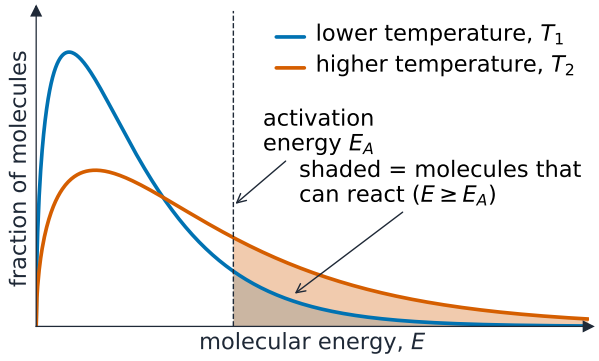
**Temperature is average KE**

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

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

# 5 Solutions

# Solutions and solubility

A **solution** 溶液: **solute** 溶质 dissolved in **solvent** 溶剂.

## Molarity and dilution

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

"Like dissolves like": polar in polar, nonpolar in nonpolar.

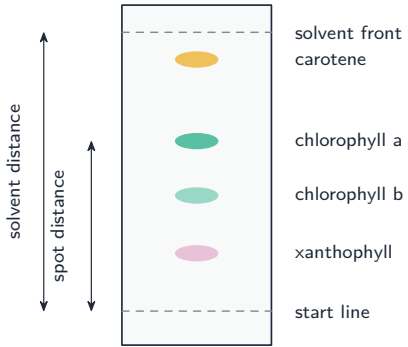
*Solute and solvent make a solution*

# 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** 解离 into separate ions, each surrounded by solvent molecules |
| **the orientation** | water's $\delta-$ oxygen faces the cations, its $\delta+$ 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



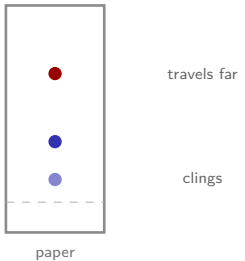

Paper chromatography separates a mixture as the solvent rises up the paper

# 6 Separation

# Separation methods

Mixtures separate by physical differences:

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



# 7 Spectroscopy

# Spectroscopy and the EM spectrum

Absorbed light matches an energy gap:

**Photon energy**

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

- microwave – **rotational** 转动
- **infrared** 红外 – **vibrational** 振动
- UV/visible – **electronic** 电子跃迁


$\lambda \Rightarrow$ higher energy.

# 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.

# The Beer-Lambert law

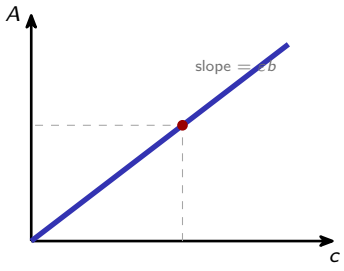
A coloured solution's **absorbance** 吸光度 is proportional to its concentration:

## Beer-Lambert

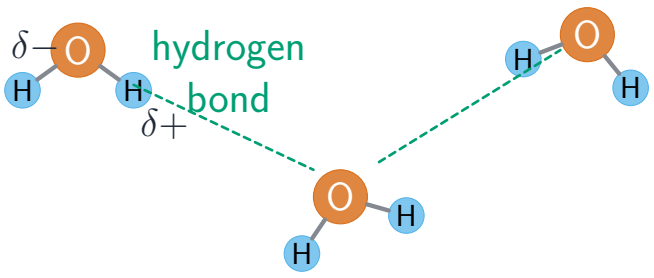
$$A = \varepsilon bc$$

$\varepsilon$ = **molar absorptivity** 摩尔吸光系数, $b$ = **path length** 光程.

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



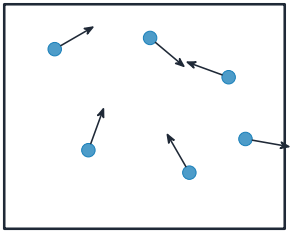
# Intermolecular and Interparticle Forces



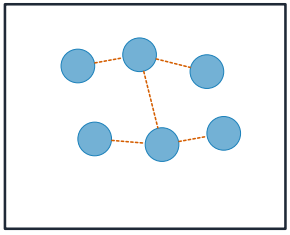

O is $\delta-$ (lone pairs),
H is $\delta+$

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

# The Ideal Gas Law



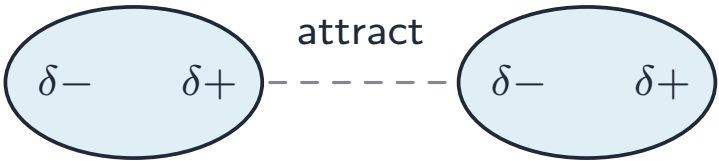
**ideal gas** (model)
point particles, *no* forces



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

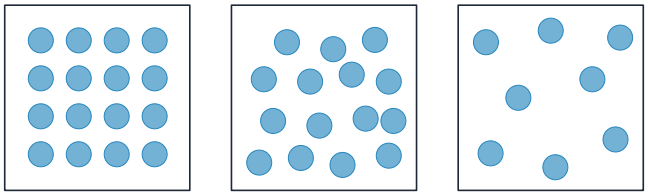
An ideal gas is a model of point particles with no forces between them

# 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 $M_1V_1 = M_2V_2$.
- "Like dissolves like" – polar/ionic solutes dissolve in polar solvents, non-polar in non-polar.

# 8 Recap

# 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**

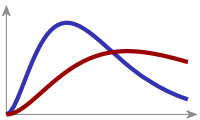
$M = n/V$; dilution $M_1V_1 = M_2V_2$; like dissolves like

**4 Light**

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

# Check yourself

1. Which IMF gives water its unusually high boiling point?
2. At the same temperature, does He or $O_2$ 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