Introduction to the particulate nature of matter

International Baccalaureate · IB Diploma · Chemistry · SL

Scope and prerequisites

This handout teaches Structure 1.1, Introduction to the particulate nature of matter, for the course first assessed in 2025. All of it is SL and HL content.

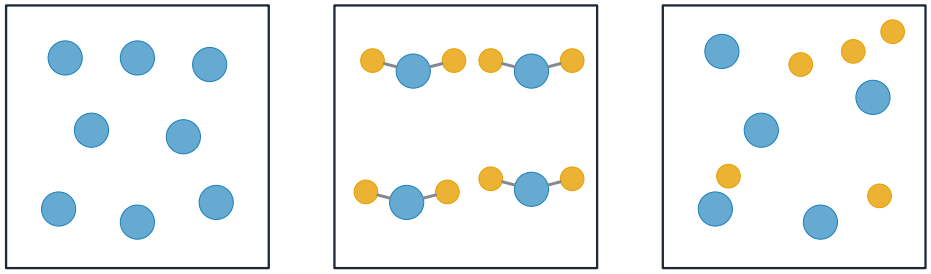
Prerequisites: the idea that matter is made of particles, and reading a temperature–time graph. This is the first unit of the course; the rest of the course builds on it.

Elements, compounds and mixtures

Why is sea water not a pure substance?

Sea water tastes of salt, but heat it in a pan and pure water leaves as steam while the salt stays behind. Sea water is a mixture: water and salt are mixed but not chemically joined, so a physical process can separate them. Water itself cannot be split by boiling; it needs a chemical reaction, such as electrolysis, to give hydrogen and oxygen.

Three kinds of matter



element
only one type of atom

compound
different atoms bonded in a fixed ratio

mixture
not bonded, easy to separate

Elements and compounds are pure substances; a mixture is not

- An **element** contains only one type of atom, for example iron, Fe, or oxygen, O_2.
- A **compound** contains two or more elements **chemically bonded** in a fixed ratio, for example water, H_2O. Its properties differ from those of its elements: sodium and chlorine are dangerous, sodium chloride is table salt.
- A **mixture** contains two or more substances that are not chemically bonded. Its composition can vary, and each part keeps its own properties.

Homogeneous and heterogeneous mixtures

- A **homogeneous** mixture has the same composition throughout, and you cannot see the parts: salt water, air, an alloy such as brass.
- A **heterogeneous** mixture does not have a uniform composition, and the parts can often be seen: sand in water, oil and water, granite.

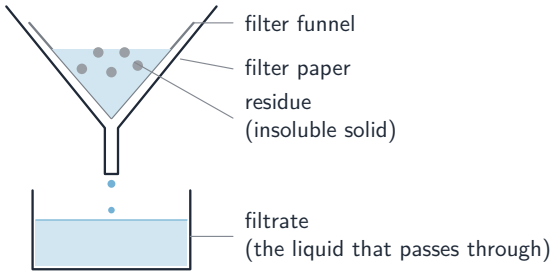
Separating mixtures

Choosing a method

Mixtures are separated by physical methods that use a difference in physical properties: solubility, boiling point, or particle size.

Mixture	Method	Property used
insoluble solid in a liquid (sand and water)	**filtration**	particle size
dissolved solid in a liquid (salt water), solid wanted	**evaporation** or **crystallization**	the solvent evaporates, the solute does not
dissolved solid in a liquid, liquid wanted	simple **distillation**	different boiling points
two miscible liquids (ethanol and water)	fractional distillation	different boiling points
two soluble solids (sand and salt)	**solvation** (dissolve), then filter and evaporate	different solubility
coloured substances in a solution (inks, dyes)	paper chromatography (Structure 2.2)	attraction to two phases

Filtration



The solid that stays in the paper is the residue; the liquid that passes through is the filtrate

The filter paper has tiny holes. Dissolved particles and the solvent pass through; larger, undissolved particles do not.

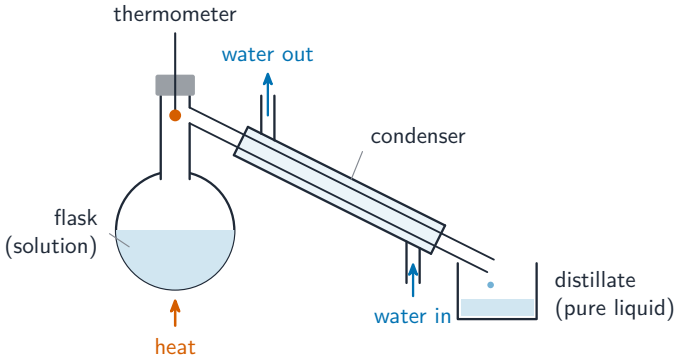
Crystallization

To get crystals of a soluble solid, such as copper(II) sulfate, from its solution:

1. Heat the solution to evaporate some of the water until it is **saturated** : a drop on a cold glass rod forms crystals.
2. Leave it to cool slowly. The solid is less soluble when cold, so crystals form.
3. Filter off the crystals, wash them with a little cold water and dry them.

Evaporating all the water instead gives a powder, and can decompose some solids or remove their water of crystallization.

Distillation



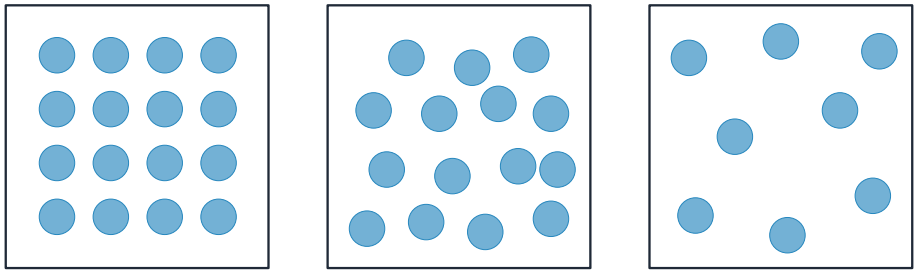
The liquid boils off, condenses in the cold condenser and collects as the distillate

- The liquid with the lower boiling point boils first. The vapour cools in the condenser and turns back into liquid.
- Cold water enters the condenser at the bottom and leaves at the top, so the jacket stays full.
- The thermometer bulb sits level with the side arm, so it reads the temperature of the vapour that is leaving.
- In fractional distillation a column filled with glass beads separates liquids whose boiling points are close: crude oil, or ethanol (78 °C) from water (100 °C).

The kinetic molecular theory

Particles in the three states

The **kinetic molecular theory** says that all matter is made of particles that are always moving. How the particles are arranged and how they move decides the state.



solid
regular, touching
vibrate in place

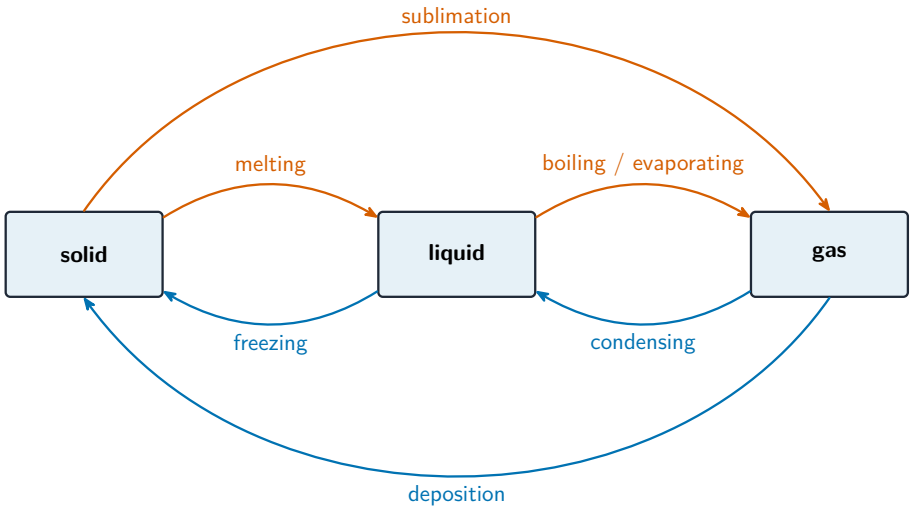
liquid
close, random
slide past each other

gas
far apart, random
move fast

The state depends on how close the particles are and how freely they move

	Solid	Liquid	Gas
Arrangement	regular, close together	random, close together	random, far apart
Movement	vibrate about fixed positions	move past each other	move fast in all directions
Forces between particles	strong	weaker	very weak
Shape	fixed	takes the shape of the container	fills the container
Volume	fixed	fixed	fills the container
Compressible?	no	hardly	yes

Changes of state



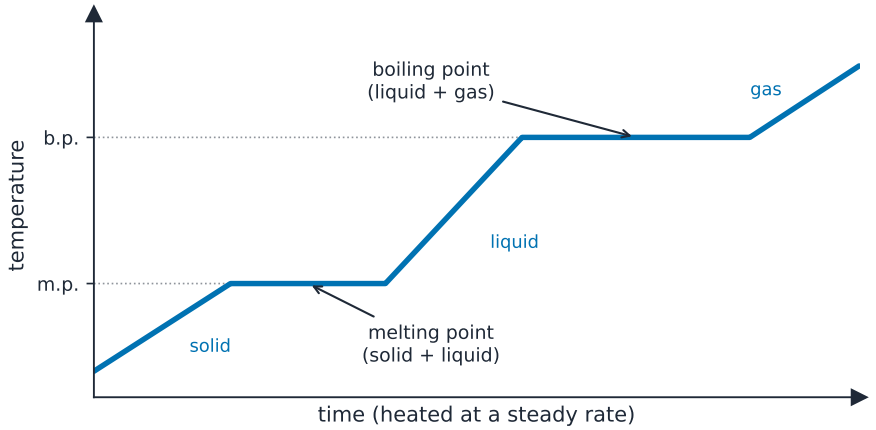
orange: particles gain energy blue: particles lose energy

Orange changes need energy; blue changes release it

- Melting, boiling or evaporating, and **sublimation** (solid straight to gas) are **endothermic** : energy is taken in to overcome the forces between particles.
- Freezing, condensing and **deposition** (gas straight to solid) are **exothermic** : energy is released as forces form.
- Dry ice (solid CO_2) and iodine sublime at atmospheric pressure. Frost forming on a cold window is deposition.

Evaporation and boiling are different. Evaporation happens only at the surface and at any temperature: the fastest particles escape. Boiling happens throughout the liquid, at one temperature, the boiling point, and forms bubbles.

■ The heating curve



plateau: attractions overcome; temperature constant

On each flat part the energy separates particles instead of making them move faster

- On the sloping parts the energy increases the kinetic energy of the particles, so the temperature rises.
- On the flat parts (the melting point and the boiling point) the energy is used to overcome the attractions between particles. The kinetic energy, and so the temperature, stays the same.
- The boiling plateau is longer than the melting plateau: it takes much more energy to separate the particles completely than to let them slide.
- A pure substance melts and boils at a sharp, fixed temperature. An impure substance melts over a range of temperatures, below the pure melting point. This is a simple test of **purity** .

Temperature and kinetic energy

■ The kelvin scale

The temperature of a substance is a measure of the **average kinetic energy** of its particles. In kelvin (K), the temperature is directly proportional to the average kinetic energy: double the temperature in K and you double the average kinetic energy.

$$T/\mathrm{K} = \theta/^\circ\mathrm{C} + 273$$

- 0 K, **absolute zero** (−273 °C), is the lowest possible temperature: the particles have the minimum possible kinetic energy.
- There are no negative temperatures on the kelvin scale, and no degree sign: write 298 K, not 298 °K.

- A temperature change of 1 K is the same as a change of 1 °C.

> **Worked example: converting temperatures**
>
> - 25 °C: $25 + 273 = 298$ K.
> - 77 K (liquid nitrogen boils): $77 - 273 = -196$ °C.
> - Warming a gas from 27 °C to 327 °C, the temperature goes from 300 K to 600 K, so the average kinetic energy doubles. In degrees Celsius it looks like a twelve-fold rise, which is wrong.

Practice

1. Classify each as an element, a compound or a mixture: (a) carbon dioxide, (b) air, (c) copper, (d) sea water, (e) sodium chloride.
2. Classify each mixture as homogeneous or heterogeneous: (a) vinegar, (b) muddy water, (c) brass, (d) a fruit salad.
3. Describe how to get pure, dry crystals of copper(II) sulfate from a mixture of the solid with sand.
4. Explain why the thermometer bulb in a distillation is placed at the side arm, not in the liquid.
5. Explain how water can be collected from salty water by distillation.
6. Describe the arrangement and movement of particles in a liquid.
7. Explain why gases are easy to compress but liquids are not.
8. Name the change of state when (a) solid iodine turns to a purple vapour, (b) water vapour forms ice on a freezer wall.
9. Explain why the temperature stays constant while ice melts, although it is being heated.
10. A sample melts between 112 °C and 117 °C. The pure substance melts at 122 °C. What does this show?
11. Convert (a) −40 °C to K, (b) 373 K to °C.
12. A gas is heated from 200 K to 400 K. State how its average kinetic energy changes. Explain why heating from 20 °C to 40 °C does not double it.

Answers.

1. (a) compound, (b) mixture, (c) element, (d) mixture, (e) compound.
2. (a) homogeneous, (b) heterogeneous, (c) homogeneous, (d) heterogeneous.
3. Add water and stir (copper sulfate dissolves, sand does not). Filter: sand stays as the residue. Heat the filtrate until saturated, cool so crystals form, filter them off, wash with a little cold water and dry.
4. The thermometer must measure the temperature of the vapour that is distilling over, which tells you which substance is being collected.
5. Heat the salty water: the water boils and its vapour passes into the condenser, cools and condenses, and drips into the beaker. The salt has a very high boiling point and stays in the flask.
6. Particles are close together in a random arrangement; they move and slide past each other.
7. Gas particles are far apart, with a lot of empty space between them; liquid particles are already close together.
8. (a) sublimation, (b) deposition.
9. The energy is used to overcome the forces between particles in the solid, not to increase their kinetic energy, so the temperature does not change until all the ice has melted.
10. The sample is impure: it melts over a range, below the pure melting point.
11. (a) $-40 + 273 = 233$ K, (b) $373 - 273 = 100$ °C.
12. It doubles, because kinetic energy is proportional to the temperature in K. 20 °C to 40 °C is 293 K to 313 K, an increase of only about 7%.

Exam tips

- “Mixture” means not chemically bonded; “compound” means chemically bonded in a fixed ratio.
- Name the property a separation uses: solubility, particle size or boiling point.
- On a heating curve, say that energy overcomes the forces between particles; the kinetic energy and temperature stay constant.
- Always convert to kelvin before comparing kinetic energies or using a gas law; write K without a degree sign.
- Evaporation happens at the surface at any temperature; boiling happens throughout the liquid at the boiling point.