MEGA LECTURE

IB DP CHEMISTRY

Structure 1: Models of the Particulate Nature of Matter

Structure 1.1 Introduction to the Particulate Nature of Matter

Revision Notes · Standard and Higher Level

Fahad H. Ahmad

+92 323 509 4443 | Megalecture.com

Original notes prepared for the IB Diploma Programme Chemistry course (first assessment 2025)

What the syllabus requires

Structure 1.1 is the starting point of the whole course: it establishes the particulate (particle) model of matter that every later topic builds on. Use the checklist below as a final revision sweep. All of this content is common to SL and HL; the few HL extension ideas are flagged with **(HL)**.

Understanding	You should be able to...
States of matter	Describe the arrangement, motion and relative energy of particles in solids, liquids and gases using the kinetic molecular theory.
Changes of state	Name every change of state and explain the energy changes involved, distinguishing kinetic from potential energy during a phase change.
Temperature	Convert between Celsius and kelvin, define absolute zero, and relate temperature to the average kinetic energy of particles.
Elements, compounds, mixtures	Define and distinguish these, and classify homogeneous and heterogeneous mixtures.
Separating mixtures	Select an appropriate physical technique (filtration, evaporation, distillation, chromatography) for a given mixture.
Pure vs impure	Use melting and boiling behaviour to distinguish a pure substance from a mixture.

> **Exam note:** This topic is assessed mostly through short-answer and multiple-choice questions. Marks are won by using precise particle language and by never confusing the ideas of temperature, heat and kinetic energy.

1. The particulate model of matter

Chemistry rests on a single powerful idea: all matter is made of extremely small particles (atoms, molecules or ions) that are in constant motion and that attract one another. This is the **particulate model** (or kinetic molecular theory). It is a *model* — a simplified mental picture — but it successfully explains states of matter, changes of state, dissolving, diffusion, gas pressure and much more.

The core assumptions

- Matter is made of tiny discrete particles with empty space between them.
- The particles are in continual, random motion; the higher the temperature, the faster they move on average.
- Particles attract one another through intermolecular (or interparticle) forces; these forces are strongest when particles are close together.
- The state of a substance is decided by the balance between the kinetic energy of the particles and the strength of the forces holding them together.

Evidence for moving particles: diffusion

Diffusion is the net movement of particles from a region of higher concentration to a region of lower concentration, caused by their random motion. It is direct, everyday evidence that particles move.

- A drop of coloured dye spreads through still water without stirring.

- The smell of perfume or ammonia reaches the far side of a room as gas particles mix with the air.
- Diffusion is fastest in gases (particles move fast and are far apart), slower in liquids, and effectively negligible in solids.

Because lighter particles travel faster at a given temperature, a gas of low molar mass diffuses more quickly than a heavier gas — ammonia (NH_3) diffuses faster than hydrogen chloride (HCl), for instance. **Brownian motion**, the jerky random movement of smoke or pollen particles seen under a microscope, is further evidence: the visible particles are being buffeted by fast-moving, invisible gas or liquid particles.

> **Key point:** Diffusion and Brownian motion are the standard experimental evidence examiners expect you to cite for the constant, random motion of particles.

2. The three states of matter

Solids, liquids and gases differ in how their particles are arranged, how they move and how much energy they have. The differences all follow from the balance between interparticle forces and particle kinetic energy.

Property	Solid	Liquid	Gas
Arrangement	Regular, closely packed, ordered lattice	Close together but irregular (disordered)	Far apart, random, mostly empty space
Movement	Vibrate about fixed positions only	Move around and slide past one another	Move rapidly in all directions
Forces between particles	Very strong	Moderate	Very weak (almost none)
Relative kinetic energy	Lowest	Intermediate	Highest
Shape	Fixed	Takes shape of container	Fills container
Volume	Fixed	Fixed	Not fixed (expands to fill)
Compressibility	Almost none	Almost none	Easily compressed
Density	Usually highest	Intermediate	Very low

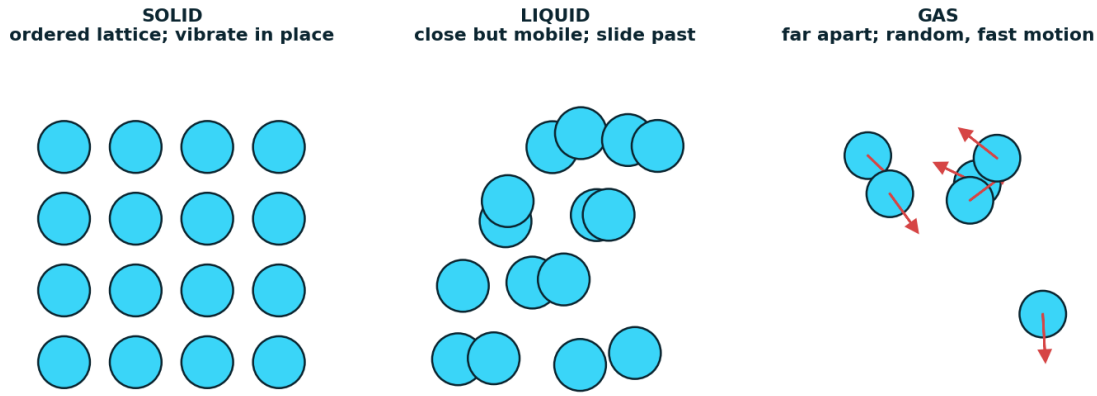

Figure 1. Kinetic molecular model of the three states. Particles form an ordered, closely packed lattice and only vibrate in a solid; are close together but disordered and mobile in a liquid; and are far apart, random and fast-moving in a gas.

Notation and everyday examples

In equations, the physical state is shown by a **state symbol**: (s) solid, (l) liquid, (g) gas and (aq) aqueous (dissolved in water). For example, water can exist as H_2O(s) ice, H_2O(l) liquid water and H_2O(g) steam — the same particles in three different states.

> **Common misconception:** When a solid is heated the particles do **not** get bigger and they do **not** gain more particles. They gain energy, vibrate more and move slightly further apart, so the *substance* expands — but each individual particle is unchanged in size.

3. Changes of state

A change of state is a physical change: no new substances are made and the change is reversible. Adding energy pushes a substance toward the gas state; removing energy pushes it toward the solid state.

Change of state	From → to	Energy	Everyday example
Melting (fusion)	solid → liquid	absorbed	ice → water
Freezing (solidifying)	liquid → solid	released	water → ice
Boiling / evaporation (vaporisation)	liquid → gas	absorbed	water → steam
Condensation	gas → liquid	released	steam → water
Sublimation	solid → gas	absorbed	solid CO_2 (dry ice), iodine
Deposition	gas → solid	released	water vapour → frost

Boiling versus evaporation

Both convert liquid to gas. **Evaporation** happens only at the surface, at any temperature below the boiling point, because a few high-energy particles at the surface escape. **Boiling** happens throughout the liquid at a fixed temperature (the boiling point) where bubbles of vapour form in the bulk of the liquid.

Heating and cooling curves

If a solid is heated steadily and its temperature is plotted against time, the graph — a **heating curve** — shows sloping sections (a single state warming) separated by two flat plateaus (melting, then boiling). A cooling curve is the mirror image. Interpreting these curves is a favourite exam skill.

Region of heating curve	What is happening	Temperature	Energy goes into...
Sloping (solid)	Solid warms up	rising	kinetic energy of particles
First plateau	Melting: solid → liquid coexist	constant (m.p.)	potential energy (breaking forces)
Sloping (liquid)	Liquid warms up	rising	kinetic energy of particles
Second plateau	Boiling: liquid → gas coexist	constant (b.p.)	potential energy (separating particles)
Sloping (gas)	Gas warms up	rising	kinetic energy of particles

Kinetic energy versus potential energy — the key idea

During a change of state the temperature stays **constant** even though energy is still being supplied. This is the point examiners test most often, so learn the reasoning precisely:

- Temperature is a measure of the **average kinetic energy** of the particles. While a substance is melting or boiling, the temperature is constant, so the average kinetic energy is constant.
- The energy supplied during the plateau does not speed the particles up. Instead it increases their **potential energy** by overcoming the attractive forces between them (partly on melting, fully on boiling).
- Once the change of state is complete, further heating again raises the kinetic energy and the temperature rises once more.

More energy is absorbed on boiling than on melting, because boiling must separate the particles completely (all interparticle forces broken), whereas melting only loosens the rigid lattice. This is why the boiling plateau is longer than the melting plateau on a heating curve.

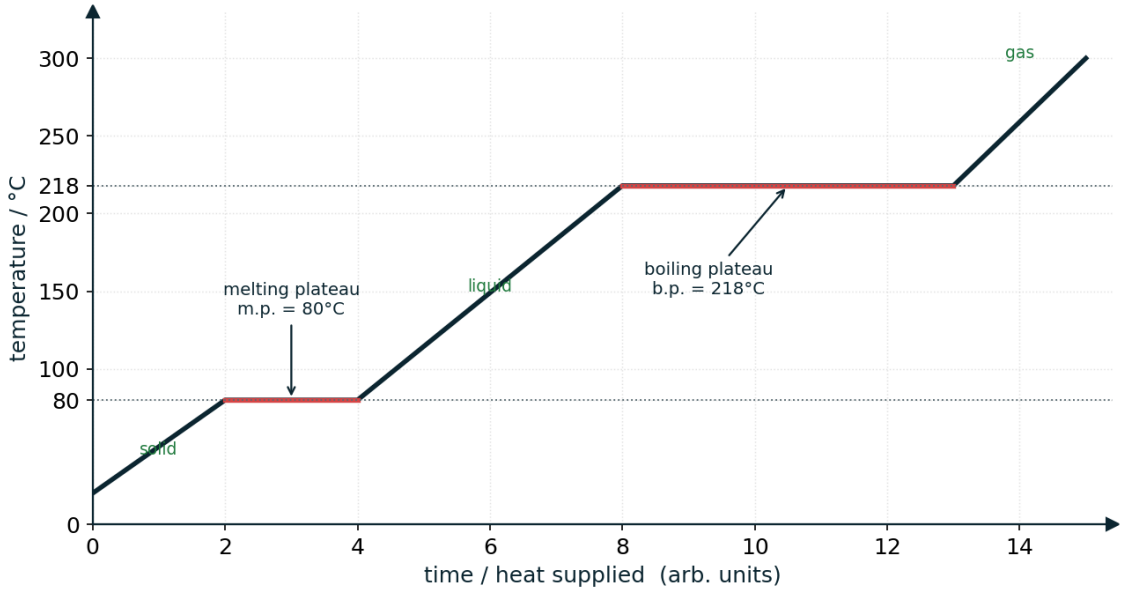

Figure 2. Heating curve of a pure substance. Temperature rises while a single state warms and is constant on the two plateaus, melting (m.p. = 80 °C) and boiling (b.p. = 218 °C), where the supplied energy raises potential, not kinetic, energy. The boiling plateau is the longer of the two.

Worked example 1 — reading a heating curve

A pure solid is heated at a constant rate. Its temperature rises, then stays at 80 °C for a while, rises again, then stays at 218 °C, and finally rises once more. (a) What is the melting point? (b) What is the boiling point? (c) During the plateau at 80 °C, why does the temperature stay constant although heating continues?

(a) The **first** plateau is melting, so the melting point is **80 °C**.

(b) The **second** plateau is boiling, so the boiling point is **218 °C**.

(c) The supplied energy is used to increase the particles' **potential energy** by overcoming the forces between them (melting the lattice), not their kinetic energy. Because the average kinetic energy is unchanged, the temperature stays constant.

4. Temperature and the kelvin scale

Temperature is a measure of the average kinetic energy of the particles in a substance. It is not the same as heat: heat is the total thermal energy transferred, whereas temperature reflects the *average* energy per particle. A spark and a bathtub of warm water can be at very different temperatures yet contain very different amounts of thermal energy.

Two temperature scales

The Celsius scale (°C) is defined around water: 0 °C is the freezing point and 100 °C the boiling point at standard pressure. The **kelvin** scale (K) is the SI scale used throughout chemistry. Its zero is **absolute zero**, the coldest possible temperature, at which particles have the minimum possible kinetic energy. A change of one kelvin equals a change of one degree Celsius, so the scales are simply shifted:

$$T\,(\mathrm{K}) = T\,(^{\circ}\mathrm{C}) + 273$$

(The precise offset is 273.15; the syllabus and data booklet use 273.) Absolute zero is therefore 0 K = −273 °C. Because temperature is proportional to average kinetic energy, and kinetic energy cannot be negative, the kelvin scale can never be negative — a key reason chemists prefer it.

Description	Celsius	Kelvin
Absolute zero	−273 °C	0 K
Freezing point of water	0 °C	273 K
Room temperature (approx.)	25 °C	298 K
Boiling point of water	100 °C	373 K

Average kinetic energy is proportional to the temperature **in kelvin**: doubling the kelvin temperature doubles the average kinetic energy of the particles. This is why gas-law and energy calculations must always use kelvin, never Celsius.

Worked example 2 — temperature conversions

(a) Convert 37 °C (body temperature) to kelvin. (b) Convert 350 K to °C. (c) A gas is heated from 27 °C to 327 °C. By what factor does the average kinetic energy of its particles increase?

(a) T = 37 + 273 = **310 K**.

(b) T = 350 − 273 = **77 °C**.

(c) In kelvin: 27 °C = 300 K and 327 °C = 600 K. Average kinetic energy is proportional to T (K), so it increases by a factor of 600/300 = **2**. (Working in Celsius here would give the wrong factor — a classic trap.)

A range of energies: the Maxwell–Boltzmann distribution

At any instant the particles in a sample do **not** all have the same kinetic energy. Because they collide and exchange energy constantly, at a given temperature there is a whole **range** (a spread) of kinetic energies: a few particles move very slowly, a few move very fast, and most have energies near the average. This spread is described qualitatively by the **Maxwell–Boltzmann distribution**.

- The curve starts at the origin (no particle has exactly zero energy), rises to a peak at the most probable energy, then falls with a long tail toward high energies.
- The **average** kinetic energy corresponds to the temperature; the peak lies just below the average because of the long high-energy tail.
- Raising the temperature shifts the whole distribution to higher energies: the peak lowers and moves right, and the fraction of high-energy particles grows. This idea returns in Reactivity 2.2 (rates and activation energy).

The existence of a range of energies also explains evaporation below the boiling point: only the fastest-moving surface particles, in the high-energy tail, have enough energy to escape the liquid.

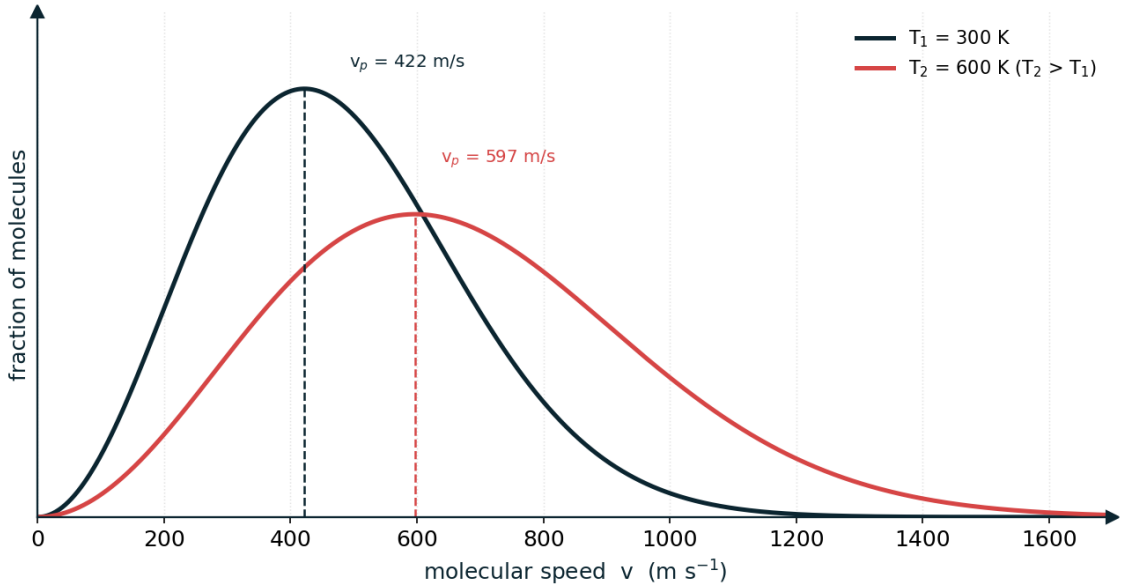

Figure 3. Maxwell–Boltzmann distribution of molecular speeds for N_2 at T_1 = 300 K and T_2 = 600 K. Raising the temperature shifts the peak to the right and broadens it. The most probable speed $v_p = \sqrt{(2kT/m)}$ increases from 422 to 597 m s^{-1}.

5. Elements, compounds and mixtures

All matter can be classified by what kinds of particles it contains and how they are combined. First distinguish **pure substances** (elements and compounds, which have a fixed composition) from **mixtures** (two or more substances physically combined in any proportion).

Type	Definition	Made of	Example
Element	A pure substance that cannot be broken down into simpler substances by chemical means; contains only one type of atom.	one kind of atom	copper (Cu), oxygen (O_2)
Compound	A pure substance in which two or more elements are chemically bonded in a fixed ratio.	atoms of different elements bonded together	water (H_2O), carbon dioxide (CO_2)
Mixture	Two or more substances physically combined and not chemically bonded; composition can vary.	elements and/or compounds not bonded	air, seawater, brass

Compounds versus mixtures

The distinction is a classic exam theme. In a **compound** the elements are chemically bonded, so it has a fixed composition, definite properties and can only be separated by a chemical reaction, and its properties differ completely from those of its elements. In a **mixture** the substances keep their own properties, the composition can vary, and they can be separated by physical means.

Feature	Compound	Mixture
Combination	Chemically bonded	Physically mixed
Composition	Fixed ratio	Variable

Feature	Compound	Mixture
Separation	Chemical reaction (e.g. electrolysis)	Physical method (e.g. filtration)
Properties	New properties, unlike the elements	Components keep their own properties
Energy change on forming	Usually a detectable energy change	Little or no energy change

> **Illustration:** Iron filings mixed with sulfur powder is a **mixture**: a magnet still removes the iron, and the ratio can be anything. Heat the mixture and it reacts to form iron(II) sulfide, FeS — a **compound** with a fixed 1:1 ratio and new properties, from which a magnet no longer removes any iron.

Homogeneous and heterogeneous mixtures

Mixtures are further classified by how uniform they are:

- A **homogeneous** mixture has a uniform composition throughout, with no visible boundaries between components — for example a salt solution, air, or an alloy such as brass. Solutions are homogeneous.
- A **heterogeneous** mixture has a non-uniform composition, with visibly distinct regions or phases — for example sand in water, oil and water, or a rock such as granite.

6. Pure substances versus mixtures: melting and boiling behaviour

A powerful practical test distinguishes a pure substance from a mixture: **a pure substance melts and boils at a single, sharp, fixed temperature, whereas a mixture melts and boils over a range of temperatures.**

Behaviour	Pure substance	Mixture
Melting point	Sharp, fixed value	Melts over a range; range is lower and broader
Boiling point	Sharp, fixed value	Boils over a range of temperatures
Heating curve plateau	Perfectly flat (horizontal)	Sloping or ill-defined during the change

This is why chemists judge purity by measuring a melting point: a sharp melting point at the expected value indicates a pure sample, while a depressed and widened melting range indicates an impurity. An impurity **lowers** the melting point and **raises** the boiling point of a substance — the principle behind salting icy roads (lowering the freezing point of water) and adding antifreeze to car radiators.

> **Exam tip:** If a question shows a heating curve with a slanted rather than flat plateau, or gives a melting *range* instead of a single value, the sample is a mixture (impure), not a pure substance.

7. Separating mixtures

Because the components of a mixture are not chemically bonded, they can be separated by **physical** methods that exploit a difference in a physical property — particle size, solubility, boiling point, or affinity for

a surface. Choosing the right method for a given mixture is a common exam task.

Technique	Separates...	Property exploited	Example
Filtration	An insoluble solid from a liquid	Particle size (residue vs filtrate)	Sand from water
Evaporation / crystallisation	A dissolved (soluble) solid from its solvent	The solute does not evaporate, the solvent does	Salt from salt solution
Simple distillation	A solvent (pure liquid) from a dissolved solid	Difference in boiling point; vapour is condensed	Pure water from seawater
Fractional distillation	Two or more miscible liquids	Different boiling points	Ethanol from water; crude oil
(Paper) chromatography	Soluble coloured substances (e.g. dyes, pigments)	Different solubilities / affinity for the paper	Dyes in an ink or food colouring
Using a magnet	A magnetic solid from a non-magnetic one	Magnetism	Iron from a sand mixture

How the key methods work

- **Filtration:** pour the mixture through filter paper; the insoluble solid (residue) is trapped while the liquid (filtrate) passes through.
- **Evaporation / crystallisation:** heat the solution so the solvent evaporates, leaving the dissolved solid behind. Gentle evaporation to the point of crystallisation gives well-formed crystals.
- **Distillation:** boil the mixture, then cool the vapour in a condenser so it condenses back to a pure liquid (the distillate). Simple distillation recovers the solvent; fractional distillation, using a fractionating column, separates liquids with close boiling points.
- **Chromatography:** a spot of the mixture is placed on paper and a solvent moves up through it; components that are more soluble and less strongly held travel further, separating into distinct spots.

Worked example 3 — choosing a separation technique

Suggest how you would obtain: (a) pure water from a sample of muddy, salty seawater; (b) copper(II) sulfate crystals from copper(II) sulfate solution; (c) the separate dyes present in a sample of green food colouring.

(a) First **filter** to remove the insoluble mud, then **distil** the salty filtrate: the water boils off, is condensed, and collects as pure water, leaving the dissolved salt behind.

(b) **Evaporate / crystallise:** heat the solution to drive off most of the water, then leave it to cool so that copper(II) sulfate crystals form.

(c) Use **paper chromatography:** the dyes have different solubilities and affinities for the paper, so they travel different distances and separate into distinct spots.

Worked example 4 — classifying matter and using purity

A student has three unlabelled samples: (i) pure liquid X boils sharply at 78 °C; (ii) liquid Y boils over the range 78–92 °C; (iii) a shiny solid Z that a chemical test shows contains only one type of atom. Classify each sample as an element, a compound or a mixture, giving your reasoning.

(i) X boils at a single sharp temperature, so it is a **pure substance** — either an element or a compound (a single boiling point cannot alone distinguish the two).

(ii) Y boils over a **range** of temperatures, so it must be a **mixture**.

(iii) Z contains only one type of atom, so it is an **element**.

8. Common pitfalls

- Saying that particles themselves expand or grow when heated. They do not: they gain energy and move further apart, so the substance expands, but each particle is unchanged.
- Claiming the temperature rises during melting or boiling. It stays constant; the energy raises potential energy, not kinetic energy, so the temperature is unchanged.
- Confusing kinetic and potential energy. During a phase change the kinetic energy (and temperature) is constant while the potential energy changes.
- Using Celsius in energy or gas calculations. Average kinetic energy is proportional to the temperature in **kelvin**; always convert first.
- Treating heat and temperature as the same thing. Temperature is the average energy per particle; heat is total thermal energy transferred.
- Mixing up compounds and mixtures. A compound is chemically bonded with a fixed ratio; a mixture is physically combined with a variable ratio.
- Assuming a single melting or boiling point proves a substance is an element. It only proves the substance is pure; it could equally be a compound.
- Forgetting that evaporation occurs below the boiling point (surface only), whereas boiling occurs throughout the liquid at a fixed temperature.

9. Quick reference

Idea	Statement
Particulate model	All matter is made of tiny particles in constant random motion that attract one another.
States	Solid (ordered, vibrating) → liquid (close, mobile) → gas (far apart, fast).
Changes of state	melt / freeze; boil (vaporise) / condense; sublime / deposit.
Phase-change energy	Temperature (and average KE) constant; potential energy changes as forces are overcome.
Temperature scale	T (K) = T (°C) + 273; absolute zero = 0 K = −273 °C.
Temperature meaning	Proportional to average kinetic energy (in kelvin); particles have a range of energies.

Idea	Statement
Element / compound / mixture	one atom type / chemically bonded fixed ratio / physically mixed variable ratio.
Purity test	Pure = sharp fixed m.p./b.p.; mixture = melts and boils over a range.
Separation	filtration, evaporation/crystallisation, distillation, chromatography.

10. Test yourself

Attempt these without notes, then check against the worked answers below.

1. Describe the arrangement, motion and relative energy of the particles in a liquid.

2. Name the change of state in each case and state whether energy is absorbed or released: (a) frost forming on a cold window; (b) a puddle drying up; (c) solid iodine turning directly into purple vapour.

3. Convert (a) 100 °C to kelvin, (b) 200 K to °C, and (c) state the value of absolute zero in °C.

4. Explain, in terms of kinetic and potential energy, why the temperature of pure water stays at 100 °C while it is boiling, even though heat is still supplied.

5. Classify each as element, compound or mixture: (a) O_2; (b) CO_2; (c) air; (d) brass.

6. A sample melts sharply at 53 °C; a second sample of the same substance melts over 47–51 °C. What does each result tell you about purity, and why does the impure sample behave this way?

7. Describe how you would obtain pure salt (sodium chloride) and pure water separately from a sample of rock salt (salt mixed with insoluble sand).

8. State two pieces of experimental evidence for the idea that particles are in constant motion, and explain what each shows.

Answers

1. In a liquid the particles are close together but arranged irregularly (disordered). They have enough energy to move around and slide past one another, so a liquid flows and takes the shape of its container. Their kinetic energy is greater than in a solid but less than in a gas.

2. (a) Deposition (gas → solid), energy released. (b) Evaporation (liquid → gas), energy absorbed. (c) Sublimation (solid → gas), energy absorbed.

3. (a) 100 + 273 = **373 K**. (b) 200 − 273 = **−73 °C**. (c) Absolute zero = **−273 °C** (0 K).

4. Temperature measures the average kinetic energy of the particles. While the water boils, the supplied energy is used to increase the particles' potential energy by overcoming the intermolecular forces and separating the particles into a gas, not to increase their kinetic energy. Because the average kinetic energy is unchanged, the temperature stays constant at 100 °C until all the liquid has boiled.

5. (a) Element (one type of atom). (b) Compound (two elements chemically bonded in a fixed ratio). (c) Mixture (homogeneous mixture of gases). (d) Mixture (an alloy of copper and zinc).

6. The sharp melting point at 53 °C shows the first sample is pure. The second sample melts over a range and at a lower temperature, showing it is impure: an impurity lowers and broadens the melting point, so the substance no longer melts at a single sharp value.

7. Add water to dissolve the salt, then **filter** to remove the insoluble sand (residue). Take the salt solution (filtrate) and **distil** it: the water boils off and is condensed and collected as pure water, while the salt is left behind. (Alternatively, evaporate/crystallise the filtrate to recover the salt if the water is not needed.)

8. *Diffusion* — e.g. a coloured dye spreading through still water, or a smell travelling across a room — shows particles move from high to low concentration by their own random motion. *Brownian motion* — the jerky random movement of visible smoke or pollen particles under a microscope — shows they are being struck by fast-moving, invisible particles, confirming constant particle motion.