

States of matter: gases, liquids and solids (AS Chemistry notes)

These notes cover the States of matter topic of the Cambridge International AS Chemistry syllabus (9701). They explain how particles behave in gases, liquids and solids, how to use the ideal gas equation with correct units, why real gases do not always behave ideally, and how the structure of a solid explains its physical properties.

1. The kinetic theory of matter

1.1 Particles in the three states

All matter is made of particles (atoms, molecules or ions) that are in constant motion. The state of a substance depends on the balance between two things: the kinetic energy of the particles, which tends to spread them out, and the attractive forces between them, which tend to hold them together.

Feature	Solid	Liquid	Gas
Arrangement	Regular, closely packed (in a crystalline solid)	Random, closely packed	Random, far apart
Movement	Vibrate about fixed positions	Move past one another, sliding and rotating	Move rapidly in straight lines between collisions
Forces between particles	Strong enough to hold particles in place	Significant, but particles can change neighbours	Very weak (negligible in an ideal gas)
Shape and volume	Fixed shape, fixed volume	Takes the shape of its container, fixed volume	Fills its container, no fixed volume
Compressibility	Almost incompressible	Almost incompressible	Easily compressed

1.2 Temperature and kinetic energy

Particles in a sample have a spread of kinetic energies. The absolute temperature of a substance is proportional to the average kinetic energy of its particles.

Because of this, gas calculations always use temperature in kelvin, never in degrees Celsius.

Converting temperature: $T / K = \theta / ^\circ C + 273$

For example, $25\text{ }^\circ\text{C} = 298\text{ K}$ and $0\text{ }^\circ\text{C} = 273\text{ K}$. (Some data use 273.15; 273 is accepted at AS Level.)

1.3 Changes of state

On melting and boiling, the energy supplied overcomes attractive forces between particles, so the temperature stays constant while the change happens.

The stronger the forces between particles, the more energy is needed to separate them, and the higher the melting point and boiling point. This single idea explains most of the property patterns in section 5.



2. The gaseous state: ideal gases

2.1 What is an ideal gas?

An ideal gas is a model: an imaginary gas that obeys the ideal gas equation exactly at all temperatures and pressures.

Assumptions of the kinetic theory for an ideal gas (learn these):

1. The gas particles are in constant, rapid, random motion.
2. The volume of the particles themselves is negligible compared with the volume of the container (the particles are treated as points).
3. There are no intermolecular forces of attraction or repulsion between the particles.
4. Collisions between particles, and between particles and the walls of the container, are perfectly elastic: no kinetic energy is lost in a collision.
5. The average kinetic energy of the particles is proportional to the absolute temperature.

In the exam, the two assumptions most often asked for are: **negligible volume of particles** and **no intermolecular forces**.

2.2 Pressure explained by kinetic theory

Gas pressure is caused by particles colliding with the walls of the container. Each collision exerts a tiny force. Pressure is the total force per unit area of wall.

- **More particles in the same volume** (larger n): more collisions per second with the walls, so the pressure rises.
- **Smaller volume at the same temperature**: particles hit the walls more often, so the pressure rises. This is Boyle's law: at constant temperature, p is inversely proportional to V .
- **Higher temperature at constant volume**: particles move faster, so they hit the walls more often and with more force, so the pressure rises.
- **Higher temperature at constant pressure**: the gas must expand so that the collision rate per unit area stays the same. This is Charles's law: at constant pressure, V is proportional to T (in kelvin).

2.3 The ideal gas equation

The ideal gas equation: $pV = nRT$

p = pressure in pascals (Pa)

V = volume in cubic metres (m^3)

n = amount of gas in moles (mol)

R = the molar gas constant, $8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ (use the value given in your data booklet)

T = temperature in kelvin (K)

Every quantity must be in SI units. Most errors in gas calculations are unit errors.

2.4 Unit handling

Quantity	Common unit given	Convert to SI
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Pressure	kPa	$\times 10^3$ (100 kPa = 1.00×10^5 Pa)
Pressure	MPa	$\times 10^6$
Volume	dm ³	$\times 10^{-3}$ (1 dm ³ = 1×10^{-3} m ³)
Volume	cm ³	$\times 10^{-6}$ (1 cm ³ = 1×10^{-6} m ³)
Temperature	°C	+ 273
Mass	g	Keep in grams when finding moles with M in g mol ⁻¹

Exam tip: write the converted values out in a line before you substitute, for example " $p = 1.01 \times 10^5$ Pa, $V = 2.50 \times 10^{-4}$ m³, $T = 300$ K". If your answer comes out a thousand or a million times too big or too small, check the volume conversion first. Give your final answer to the same number of significant figures as the data (usually 3).

Worked example 1: finding a volume

Calculate the volume occupied by 0.250 mol of oxygen at 20 °C and 150 kPa.

Step 1, convert: $n = 0.250$ mol, $T = 20 + 273 = 293$ K, $p = 150 \times 10^3 = 1.50 \times 10^5$ Pa

Step 2, rearrange: $V = nRT / p$

Step 3, substitute: $V = (0.250 \times 8.31 \times 293) / (1.50 \times 10^5) = 608.7 / 1.50 \times 10^5 = 4.06 \times 10^{-3}$ m³

Step 4, convert if asked: 4.06×10^{-3} m³ = 4.06 dm³

Worked example 2: finding a pressure

0.0500 mol of a gas is held in a 500 cm³ container at 400 K. Calculate the pressure in kPa.

$V = 500 \times 10^{-6} = 5.00 \times 10^{-4}$ m³

$p = nRT / V = (0.0500 \times 8.31 \times 400) / (5.00 \times 10^{-4}) = 166.2 / 5.00 \times 10^{-4} = 3.32 \times 10^5$ Pa = 332 kPa

2.5 Finding relative molecular mass, M_r , from gas data

Since the amount $n = m / M$, where m is the mass in grams and M is the molar mass in g mol⁻¹, the ideal gas equation can be written:

$pV = (m / M)RT$, so $M = mRT / pV$

With m in grams and p , V , T in SI units, M comes out in g mol⁻¹. The numerical value of M is the M_r .

In the classic experiment, a known mass of a volatile liquid is injected into a heated gas syringe, vaporises completely, and the vapour volume, temperature and pressure are recorded.

Worked example 3: M_r of a vapour

0.500 g of a volatile liquid is vaporised. The vapour occupies 250 cm³ at 27 °C and 100 kPa.

Calculate the M_r of the liquid.

$m = 0.500$ g, $T = 27 + 273 = 300$ K, $p = 1.00 \times 10^5$ Pa, $V = 250 \times 10^{-6} = 2.50 \times 10^{-4}$ m³

$M = mRT / pV = (0.500 \times 8.31 \times 300) / (1.00 \times 10^5 \times 2.50 \times 10^{-4}) = 1246.5 / 25.0 = 49.9$ g mol⁻¹

$M_r = 49.9$ (no units)

You can also find M_r from the density of a gas. Density $\rho = m / V$, so $M = \rho RT / p$. Take care: if the density is in g dm^{-3} , multiply by 10^3 to get g m^{-3} before using SI pressure.

Worked example 4: M_r from density

A monatomic gas has a density of 1.78 g dm^{-3} at 273 K and 101 kPa. Calculate its M_r and suggest its identity.

$$\rho = 1.78 \text{ g dm}^{-3} = 1.78 \times 10^3 \text{ g m}^{-3}, p = 1.01 \times 10^5 \text{ Pa}$$

$$M = \rho RT / p = (1780 \times 8.31 \times 273) / (1.01 \times 10^5) = 4.04 \times 10^6 / 1.01 \times 10^5 = 40.0 \text{ g mol}^{-1}$$

A monatomic gas with M_r about 40 is argon.

2.6 Gas volumes in reactions

Combining the ideal gas equation with a balanced equation lets you predict the volume of gas made in a reaction. Find the moles of the limiting reactant, use the mole ratio to find the moles of gas, then use $V = nRT / p$. At room temperature and pressure you may instead be told to use the molar gas volume from the data booklet ($24.0 \text{ dm}^3 \text{ mol}^{-1}$ at room conditions); use whichever method the question asks for.

Worked example 5: gas from a reaction

Magnesium reacts with excess hydrochloric acid: $\text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$

Calculate the volume of hydrogen, in cm^3 , made from 0.120 g of magnesium at 298 K and 101 kPa. (A_r : Mg = 24.3)

$$n(\text{Mg}) = 0.120 / 24.3 = 4.94 \times 10^{-3} \text{ mol}; \text{ mole ratio Mg : H}_2 = 1 : 1, \text{ so } n(\text{H}_2) = 4.94 \times 10^{-3} \text{ mol}$$

$$V = nRT / p = (4.94 \times 10^{-3} \times 8.31 \times 298) / (1.01 \times 10^5) = 1.21 \times 10^{-4} \text{ m}^3$$

$$V = 1.21 \times 10^{-4} \times 10^6 = 121 \text{ cm}^3$$

3. Real gases and deviation from ideal behaviour

3.1 Why real gases are not ideal

Real gas particles do have a volume of their own, and there are always some intermolecular forces between them (at least instantaneous dipole-induced dipole forces). The ideal gas model ignores both. Whether this matters depends on the conditions.

3.2 Conditions that cause deviation

High pressure. At high pressure the particles are squeezed close together. The volume of the particles is then a significant fraction of the total volume, so the space available for movement is less than the container volume. Also, close together, the attractive forces between particles become significant.

Low temperature. At low temperature the particles move more slowly and have less kinetic energy. When two particles approach, the attractive forces have more time and more effect, so particles are pulled towards each other. Collisions with the walls are less forceful and the measured pressure is lower than the ideal gas equation predicts.

Learn this: A real gas behaves most like an ideal gas at **high temperature** and **low pressure**. Under these conditions the particles are far apart (so their own volume is negligible) and moving fast (so intermolecular forces have little effect).

3.3 Which gases deviate most?

The stronger the intermolecular forces, and the larger the particles, the greater the deviation from ideal behaviour.

- Helium, H_2 and N_2 are close to ideal: small, non-polar particles with only weak instantaneous dipole-induced dipole forces.
- Polar molecules such as HCl , and molecules that form hydrogen bonds such as NH_3 and H_2O , deviate much more, because their intermolecular forces are stronger.
- Large molecules with many electrons, such as butane, also deviate more, because their instantaneous dipole-induced dipole forces are stronger and the molecules occupy more space.

3.4 Showing deviation on a graph

For one mole of an ideal gas, $pV / RT = 1$ at every pressure. A graph of pV / RT against p for an ideal gas is therefore a horizontal line at 1. For real gases the line curves away from 1. At moderate pressures attractive forces usually make pV / RT fall below 1. At very high pressures the volume of the particles dominates and pV / RT rises above 1. The further the line is from 1, the less ideal the gas.

Exam tip: when you explain deviation, always link the condition to a specific assumption. "At high pressure the particles are close together, so the volume of the particles is no longer negligible compared with the volume of the container" earns more credit than "the particles are squashed".

4. The liquid state and vapour pressure

4.1 Evaporation

In a liquid, particles have a range of kinetic energies. At any temperature, some particles at the surface have enough energy to overcome the attractive forces of their neighbours and escape into the gas phase. This is evaporation. It happens at all temperatures, not just at the boiling point.

4.2 Vapour pressure

If a liquid is placed in a closed container, particles evaporate into the space above it. Some of these vapour particles hit the liquid surface and condense. Eventually the rate of evaporation equals the rate of condensation and a dynamic equilibrium is set up:

liquid $\rightleftharpoons$ vapour

Definition: The vapour pressure of a liquid is the pressure exerted by its vapour when the vapour is in dynamic equilibrium with the liquid in a closed container at a given temperature.

- Vapour pressure increases with temperature, because more particles have enough energy to escape.
- At a given temperature, a liquid with weaker intermolecular forces has a higher vapour pressure. Such a liquid is described as volatile. For example, propanone has a higher vapour pressure than water at room temperature, because water molecules are held by hydrogen bonds.
- Vapour pressure does not depend on the amount of liquid or the size of the container, as long as some liquid remains at equilibrium.

4.3 Boiling

A liquid boils when its vapour pressure equals the external pressure acting on its surface. At this point bubbles of vapour can form throughout the liquid, not just at the surface.

- The normal boiling point is the temperature at which the vapour pressure equals standard atmospheric pressure (101 kPa).
- At high altitude the external pressure is lower, so the vapour pressure reaches it at a lower temperature: water boils below 100 °C.

5. The solid state: lattice structures

5.1 Crystalline solids and lattices

Definition: A lattice is a regular, repeating three-dimensional arrangement of particles (atoms, ions or molecules) in a crystalline solid.

5.2 Giant ionic lattices

A giant ionic lattice is a regular arrangement of positive and negative ions held together by strong electrostatic attractions between oppositely charged ions in all directions. Sodium chloride is the standard example: each Na^+ ion is surrounded by six Cl^- ions, and each Cl^- ion by six Na^+ ions (6 : 6 coordination). Magnesium oxide has the same arrangement.

- **High melting and boiling points:** a lot of energy is needed to overcome the many strong electrostatic attractions throughout the lattice.
- **Melting point depends on charge and size:** MgO (Mg^{2+} , O^{2-}) melts at a much higher temperature than NaCl (Na^+ , Cl^-), because the ions have higher charges and are smaller, so the attractions are stronger.
- **Electrical conductivity:** solids do not conduct, because the ions are fixed in the lattice. When molten or dissolved in water, the ions are free to move and carry charge, so they conduct.
- **Brittle:** if a force shifts one layer of ions, ions of the same charge come next to each other and repel, so the crystal splits.

5.3 Giant covalent (giant molecular) lattices

In a giant covalent lattice, a very large number of atoms are joined by strong covalent bonds in a continuous network. There are no separate molecules.

Diamond: each carbon atom is covalently bonded to four others in a tetrahedral arrangement (bond angle 109.5°), forming a rigid three-dimensional network.

- Very high melting point: many strong covalent bonds must be broken.
- Extremely hard: the strong bonds act in all directions, so the structure resists distortion. Used in cutting tools.
- Does not conduct electricity: all four outer electrons of each carbon are used in bonding, so there are no delocalised electrons.

Graphite: each carbon atom is covalently bonded to three others in flat layers of hexagonal rings. The fourth outer electron of each carbon is delocalised over the layer. The layers are held together by weak

instantaneous dipole-induced dipole forces.

- Very high melting point: strong covalent bonds within the layers must be broken.
- Conducts electricity along the layers: the delocalised electrons can move.
- Soft and slippery: the weak forces between layers allow the layers to slide over each other.
Used as a lubricant and in pencils.

Silicon(IV) oxide, SiO_2 : each silicon atom is bonded to four oxygen atoms, and each oxygen atom to two silicon atoms, in a tetrahedral network. It has a high melting point, is hard, does not conduct, and is insoluble in water.

5.4 Metallic lattices

A metallic lattice is a regular arrangement of positive metal ions surrounded by a "sea" of delocalised electrons. Metallic bonding is the electrostatic attraction between the positive ions and the delocalised electrons. Copper is a common example.

- **High melting points (usually):** strong attraction between ions and delocalised electrons. Melting point increases with more delocalised electrons per atom and smaller ions, for example Mg higher than Na, Al higher than Mg.
- **Good electrical conductors, solid and liquid:** the delocalised electrons move through the structure when a potential difference is applied.
- **Malleable and ductile:** layers of ions can slide over each other without breaking the metallic bond, because the delocalised electrons move with them and still hold the ions together.

5.5 Simple molecular lattices

In a simple molecular lattice, small molecules are held in a regular arrangement by weak intermolecular forces. Inside each molecule the atoms are joined by strong covalent bonds, but these are not broken when the substance melts or boils.

Iodine, I_2 : I_2 molecules are held in a regular lattice by instantaneous dipole-induced dipole forces. These forces are weak, so iodine has a low melting point and sublimes on gentle warming to give a purple vapour. It does not conduct, as there are no ions or delocalised electrons.

Ice, H_2O : each water molecule forms hydrogen bonds to four neighbours in a tetrahedral arrangement. This gives an open lattice with a lot of empty space. When ice melts, some hydrogen bonds break and the molecules can pack more closely, so **ice is less dense than liquid water** and floats. Hydrogen bonds are stronger than other intermolecular forces, which is why ice melts at a higher temperature than solids of other small molecules of similar size.

- Low melting and boiling points: only weak intermolecular forces need to be overcome, not covalent bonds.
- Do not conduct electricity: molecules are neutral and there are no free electrons.

Exam tip: the most common mistake in this topic is saying that a simple molecular substance melts because "covalent bonds break". They do not. Melting and boiling of a simple molecular substance overcome intermolecular forces only. Always name the force: instantaneous dipole-induced dipole, permanent dipole-permanent dipole, or hydrogen bonding.

5.6 Summary: how structure explains properties

Structure	Particles	Forces overcome on melting	Melting point	Conducts when solid?
Giant ionic	Ions	Strong electrostatic attractions between ions	High	No
Giant covalent	Atoms	Strong covalent bonds	Very high	No (graphite yes)
Metallic	Positive ions and delocalised electrons	Metallic bonding (partly)	Usually high	Yes
Simple molecular	Molecules	Weak intermolecular forces	Low	No

Structure	Conducts when molten?	Example
Giant ionic	Yes	NaCl, MgO
Giant covalent	Not applicable (usually decompose or sublime)	Diamond, graphite, SiO ₂
Metallic	Yes	Cu, Mg, Al
Simple molecular	No	I ₂ , ice, CO ₂ (s)

5.7 Deducing structure from data

Given melting points and conductivities, ask in this order:

1. Does it conduct when solid? If yes, it is metallic (or graphite).
2. If not, does it conduct when molten or in solution? If yes, it is ionic.
3. If not, is the melting point very high? If yes, it is giant covalent.
4. If the melting point is low, it is simple molecular.

Quick check

Use $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ and $T / \text{K} = \theta / ^\circ\text{C} + 273$. Give numerical answers to 3 significant figures.

1. State two assumptions made in the kinetic theory about the particles of an ideal gas.
2. Convert each value to SI units: (a) 450 cm³ (b) 3.6 dm³ (c) 85 kPa (d) -23 °C.
3. Calculate the amount, in moles, of gas in a 2.40 dm³ flask at 101 kPa and 298 K.
4. Calculate the temperature, in °C, at which 0.100 mol of an ideal gas occupies 1.50 dm³ at a pressure of 120 kPa.
5. 0.284 g of a volatile liquid is completely vaporised in a gas syringe. The vapour occupies 98.0 cm³ at 373 K and 101 kPa. Calculate the M_r of the liquid.
6. Calculate the density of nitrogen gas, N₂, in g dm⁻³, at 298 K and 100 kPa. (A_r : N = 14.0)
7. 5.00 g of calcium carbonate is heated until it decomposes completely: $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$. Calculate the volume of carbon dioxide, in dm³, produced at 300 K and 100 kPa. (A_r : Ca = 40.1, C = 12.0, O = 16.0)
8. State the conditions of temperature and pressure under which a real gas behaves most like an

ideal gas. Explain your answer.

9. Explain why ammonia, NH_3 , deviates more from ideal behaviour than nitrogen, N_2 , under the same conditions.
10. Explain, in terms of vapour pressure, why water boils at a temperature below 100°C at the top of a high mountain.
11. Explain why magnesium oxide has a much higher melting point than sodium chloride.
12. Diamond and graphite are both forms of carbon with giant covalent structures. Explain why graphite conducts electricity but diamond does not.
13. Silicon(IV) oxide melts at a very high temperature, but carbon dioxide is a gas at room temperature. Explain this difference in terms of structure and bonding.
14. Explain why metals are malleable and why they conduct electricity.
15. Substance X has a melting point of 114°C and does not conduct electricity either as a solid or as a liquid. Deduce the type of structure of X and explain your reasoning.

1. Any two of: the particles have negligible volume compared with the volume of the container; there are no intermolecular forces between the particles; collisions are perfectly elastic; the particles are in constant random motion; average kinetic energy is proportional to the absolute temperature.

2. (a) $450 \times 10^{-6} = 4.50 \times 10^{-4} \text{ m}^3$ (b) $3.6 \times 10^{-3} \text{ m}^3$ (c) $85 \times 10^3 = 8.5 \times 10^4 \text{ Pa}$ (d) $-23 + 273 = 250 \text{ K}$

3. $p = 1.01 \times 10^5 \text{ Pa}$, $V = 2.40 \times 10^{-3} \text{ m}^3$, $T = 298 \text{ K}$

$n = pV / RT = (1.01 \times 10^5 \times 2.40 \times 10^{-3}) / (8.31 \times 298) = 242.4 / 2476 = 0.0979 \text{ mol}$

4. $p = 1.20 \times 10^5 \text{ Pa}$, $V = 1.50 \times 10^{-3} \text{ m}^3$, $n = 0.100 \text{ mol}$

$T = pV / nR = (1.20 \times 10^5 \times 1.50 \times 10^{-3}) / (0.100 \times 8.31) = 180 / 0.831 = 216.6 \text{ K} (217 \text{ K})$

In $^\circ\text{C}$: $216.6 - 273 = -56.4^\circ\text{C}$.

5. $m = 0.284 \text{ g}$, $T = 373 \text{ K}$, $p = 1.01 \times 10^5 \text{ Pa}$, $V = 98.0 \times 10^{-6} = 9.80 \times 10^{-5} \text{ m}^3$

$M = mRT / pV = (0.284 \times 8.31 \times 373) / (1.01 \times 10^5 \times 9.80 \times 10^{-5}) = 880.3 / 9.898 = 88.9 \text{ g mol}^{-1}$

$M_r = 88.9$

6. $M(\text{N}_2) = 28.0 \text{ g mol}^{-1}$. Since $\rho = m / V$ and $m / M = pV / RT$, $\rho = pM / RT$.

$\rho = (1.00 \times 10^5 \times 28.0) / (8.31 \times 298) = 2.80 \times 10^6 / 2476 = 1131 \text{ g m}^{-3}$

Divide by 10^3 to convert to g dm^{-3} : $\rho = 1.13 \text{ g dm}^{-3}$

7. $M(\text{CaCO}_3) = 40.1 + 12.0 + (3 \times 16.0) = 100.1 \text{ g mol}^{-1}$

$n(\text{CaCO}_3) = 5.00 / 100.1 = 0.0500 \text{ mol} (0.04995)$. Mole ratio $\text{CaCO}_3 : \text{CO}_2 = 1 : 1$, so $n(\text{CO}_2) = 0.0500 \text{ mol}$

$V = nRT / p = (0.04995 \times 8.31 \times 300) / (1.00 \times 10^5) = 1.25 \times 10^{-3} \text{ m}^3 = 1.25 \text{ dm}^3$

8. High temperature and low pressure. At low pressure the particles are far apart, so the volume of the particles is negligible compared with the volume of the container. At high temperature the particles have high kinetic energy and move fast, so intermolecular forces have little effect on them. Both ideal gas assumptions are then nearly true.

9. Ammonia molecules are polar and form hydrogen bonds with each other, so the intermolecular forces between NH_3 molecules are much stronger than the weak instantaneous dipole-induced dipole forces between non-polar N_2 molecules. The ideal gas model assumes no intermolecular

forces, so the assumption is much less true for ammonia, and it deviates more.

10. A liquid boils when its vapour pressure equals the external pressure. At the top of a mountain the atmospheric pressure is lower than at sea level. The vapour pressure of water increases with temperature, so it reaches this lower external pressure at a lower temperature, below 100 °C.

11. Both have giant ionic lattices. The ions in MgO (Mg^{2+} and O^{2-}) carry double the charge of the ions in NaCl (Na^+ and Cl^-), and they are smaller. So the electrostatic attractions between oppositely charged ions are much stronger in MgO, and much more energy is needed to overcome them.

12. In graphite each carbon atom forms only three covalent bonds, so one outer electron per carbon atom is delocalised across the layer. These delocalised electrons can move and carry charge along the layers. In diamond each carbon atom forms four covalent bonds, using all four outer electrons, so there are no delocalised electrons and no mobile charge carriers.

13. SiO_2 has a giant covalent structure: each Si atom is bonded to four O atoms in a continuous network, so melting requires breaking many strong covalent bonds, which needs a lot of energy. CO_2 is simple molecular: it consists of separate small molecules held together only by weak intermolecular forces (instantaneous dipole-induced dipole), which need very little energy to overcome, so it is a gas at room temperature. The covalent bonds inside CO_2 molecules are not broken on boiling.

14. Metals have a lattice of positive ions in a sea of delocalised electrons. Malleable: when a force is applied, layers of positive ions can slide over each other; the delocalised electrons move with them, so the attraction between ions and electrons is kept and the metal changes shape without breaking. Conductivity: the delocalised electrons are free to move through the lattice and carry charge when a potential difference is applied.

15. Simple molecular. It does not conduct as a solid or liquid, so it has no mobile ions or delocalised electrons: it is not ionic or metallic. Its melting point is low (114 °C), so only weak intermolecular forces are overcome on melting, which rules out a giant covalent structure. (This matches iodine.)