

Cambridge AS & A Level · Chemistry 9701 · Notes

States of matter

AS

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Read first: Atoms, molecules and stoichiometry, Chemical bonding

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What you need to know

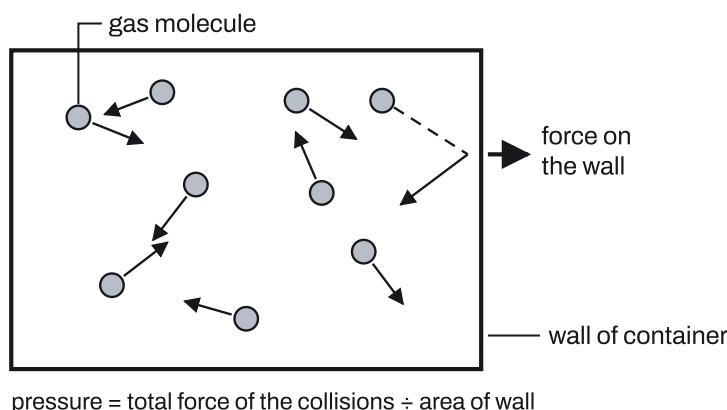
By the end of this topic you should be able to:

1. **Explain where gas pressure comes from:** gas molecules keep hitting the walls of their container, and every collision pushes on the wall.
2. **State what an ideal gas is:** its particles take up no volume, and there are no attractive forces between them.
3. **State and use the ideal gas equation** $pV = nRT$, including working out the relative molecular mass, M_r , of a gas or vapour.
4. **Describe, simply, the lattice of a crystalline solid** of each of four types:
 - giant ionic: sodium chloride and magnesium oxide;
 - simple molecular: iodine, buckminsterfullerene (C_{60}) and ice;
 - giant molecular (giant covalent): silicon(IV) oxide, graphite and diamond;
 - giant metallic: copper.
5. **Describe, explain and predict how structure and bonding affect physical properties:** melting point, boiling point, electrical conductivity and solubility.
6. **Deduce the structure and bonding of a substance** from information you are given, such as a table of properties.

Understand it

Where gas pressure comes from

The molecules of a gas move fast, in straight lines and in random directions, until they hit something. When a molecule hits the wall of its container it bounces off, and pushes on the wall as it does. Billions of collisions every second add up to a steady force on the wall. **Pressure is that force on each unit of area** (1 pascal is 1 newton on each square metre).



So anything that makes the collisions **more frequent** or **harder** raises the pressure:

- more molecules in the same space (more collisions per second);
- a smaller container (the molecules reach a wall more often);
- a higher temperature (the molecules move faster, so they hit the walls more often and harder).

Ideal gases

To get a simple equation for gases, chemists imagine an **ideal gas** whose particles:

- have **no volume** of their own (they are points), and
- have **no intermolecular forces** between them (they only interact when they collide).

The particles still have mass. They are also taken to move randomly and to bounce off each other and the walls without losing any kinetic energy overall (elastic collisions).

Real gases are not quite like this, because real molecules do take up space and do attract each other. A real gas behaves **most like an ideal gas** at:

- **low pressure**: the molecules are far apart, so their own volume is a tiny part of the container's volume, and they are too far apart to attract each other much;
- **high temperature**: the molecules move so fast that the weak attractions hardly affect them when they pass each other.

A real gas **departs most** from ideal behaviour at **high pressure** (molecules squashed together, so their own volume matters and they attract each other) and **low temperature** (slow molecules are pulled together, and the gas is close to turning into a liquid).

Between gases, the ones with **the weakest intermolecular forces** behave most ideally. Helium is the best: tiny atoms with very weak instantaneous dipole-induced dipole forces. Next come small non-polar molecules such as hydrogen and nitrogen. Large atoms (krypton), polar molecules and molecules that form **hydrogen bonds** (ammonia, water, hydrogen fluoride) behave least ideally. Water forms more hydrogen bonds per molecule than ammonia does, so steam departs from ideal behaviour even more than ammonia at the same conditions.

The ideal gas equation

For an ideal gas,

$$pV = nRT$$

where every quantity must be in **SI units**:

Symbol	Meaning	Unit	Convert from
p	pressure	Pa	kPa: $\times 1000$
V	volume	m^3	dm^3 : $\times 10^{-3}$; cm^3 : $\times 10^{-6}$
n	amount of gas	mol	$n = \frac{m}{M}$
R	molar gas constant	$8.31 \text{ J K}^{-1} \text{ mol}^{-1}$	Data Booklet
T	temperature	K	$^{\circ}\text{C}$: $+ 273$

For example, 0.0150 mol of gas at 45°C and 120 kPa occupies

$$V = \frac{nRT}{p} = \frac{0.0150 \times 8.31 \times 318}{120\,000} = 3.30 \times 10^{-4} \text{ m}^3 = 330 \text{ cm}^3.$$

The equation also explains why **one mole of any gas** takes up the same volume at the same temperature and pressure: $V = \frac{RT}{p}$ does not depend on which gas it is. The Data Booklet gives $22.4 \text{ dm}^3 \text{ mol}^{-1}$ at s.t.p. (273 K , 101 kPa) and $24.0 \text{ dm}^3 \text{ mol}^{-1}$ at room conditions.

Finding M_r with the ideal gas equation

Since $n = \frac{m}{M}$, the equation becomes $pV = \frac{m}{M}RT$, so

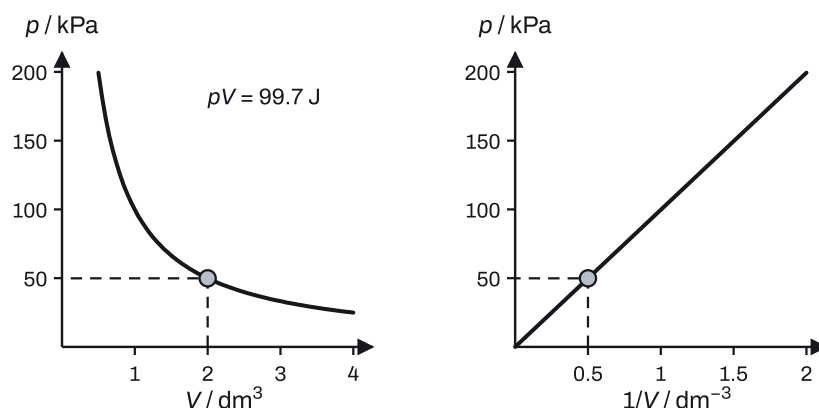
$$M = \frac{mRT}{pV}.$$

Weigh a sample, turn it into gas (for a liquid, inject it into a heated gas syringe so it all evaporates), then measure its volume, temperature and pressure. If you are given the **density** $\rho = \frac{m}{V}$ instead, then $M = \frac{\rho RT}{p}$, with ρ in g m^{-3} (a density in g dm^{-3} is multiplied by 1000).

How p , V , n and T are linked

Rearranging $pV = nRT$ shows how one quantity changes when the others are held still:

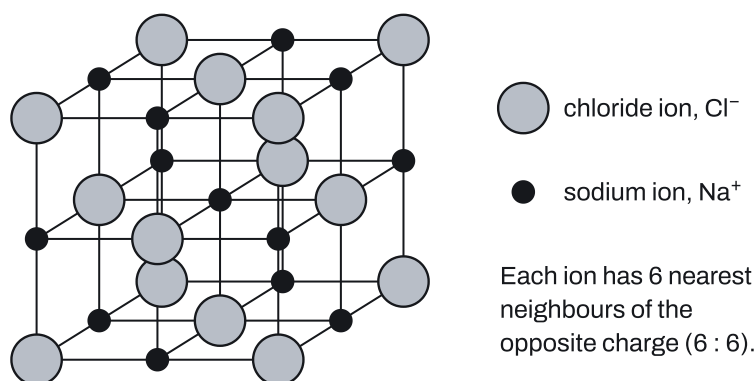
- **fixed n and T :** pV is constant, so halving the volume doubles the pressure. A graph of p against V is a curve; p against $\frac{1}{V}$ is a straight line through the origin.
- **fixed n and V :** $p \propto T$ (in kelvin).
- **fixed V and T :** $p \propto n$. A reaction that makes more moles of gas in a sealed vessel raises the pressure in proportion. In a **mixture**, each gas's share of the pressure is its share of the moles.



Four kinds of crystalline solid

In a crystalline solid the particles sit in a regular, repeating 3D arrangement called a **lattice**. What the particles are, and what holds them together, decides the properties.

Giant ionic: sodium chloride and magnesium oxide. A huge lattice of positive and negative ions held by strong electrostatic attraction between oppositely charged ions (ionic bonds) in every direction. In sodium chloride every Na^+ ion is surrounded by six Cl^- ions, and every Cl^- ion by six Na^+ ions (6 : 6). Magnesium oxide has the same arrangement with Mg^{2+} and O^{2-} ions. Its ions have double the charge, so the attraction is much stronger and its melting point is far higher (about 2850°C , against about 800°C for sodium chloride).



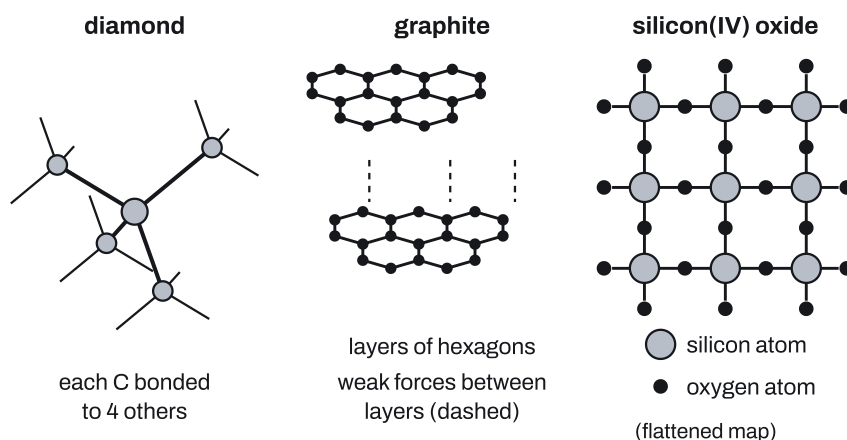
Simple molecular: iodine, buckminsterfullerene and ice. Small separate molecules, with strong covalent bonds **inside** each molecule but only weak intermolecular forces **between** the molecules.

- **Iodine:** I_2 molecules in a regular lattice, held by instantaneous dipole–induced dipole forces. It is a soft, grey-black solid that turns to a purple vapour on gentle warming.
- **Buckminsterfullerene:** C_{60} molecules, each a hollow ball of 60 carbon atoms joined in 20 hexagons and 12 pentagons, packed in a lattice and held by instantaneous dipole–induced dipole forces. Because the forces between the balls are weak, C_{60} turns to vapour at a far lower temperature than graphite or diamond, even though all three are pure carbon.
- **Ice:** H_2O molecules, each hydrogen-bonded to four others arranged tetrahedrally. This holds the molecules further apart than in liquid water, so ice is less dense than water and floats.

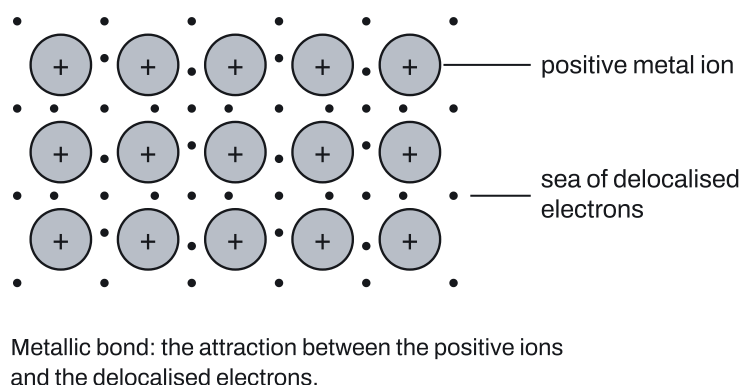
Other examples you may be asked to recognise: sulfur (S_8), white phosphorus (P_4), the halogens, and covalent non-metal compounds such as $SiCl_4$, CO_2 and $AlCl_3$.

Giant molecular (giant covalent): diamond, graphite and silicon(IV) oxide. A huge network of atoms joined by strong covalent bonds all through the structure. There are no separate molecules.

- **Diamond:** each carbon atom is covalently bonded to four others, arranged tetrahedrally. All four outer electrons are in bonds, so none is free to move.
- **Graphite:** flat **layers** of hexagons; each carbon atom is bonded to three others in its layer. Each carbon's fourth outer electron is **delocalised** along the layer, so graphite conducts electricity along the layers. The layers are held together only by weak instantaneous dipole–induced dipole forces, so they slide over each other, which makes graphite soft and slippery.
- **Silicon(IV) oxide**, SiO_2 : each silicon atom is covalently bonded to four oxygen atoms (tetrahedrally), and each oxygen atom to two silicon atoms. Hence the formula SiO_2 .



Giant metallic: copper. A regular lattice of **positive metal ions** in a "sea" of **delocalised electrons**. The metallic bond is the attraction between the positive ions and the delocalised electrons. The electrons can move through the whole structure.



Why the properties follow from the structure

Melting and boiling points. To melt a solid you must overcome whatever holds its particles in place.

- Giant ionic, giant covalent and giant metallic: many **strong** bonds must be broken or loosened throughout the lattice, which needs a lot of energy, so the melting points are **high**. Melting silicon(IV) oxide or diamond breaks covalent bonds.
- Simple molecular: only the **weak intermolecular forces** are overcome; the covalent bonds inside the molecules stay whole. So melting and boiling points are **low**. Of the solids in this topic, only the giant covalent ones break covalent bonds when they melt.
- Among simple molecular substances, bigger molecules (more electrons) have stronger instantaneous dipole–induced dipole forces, and hydrogen bonding makes the forces stronger still. That is why ice melts at 0 °C, higher than many molecules of similar size.
- Among ionic solids, higher ionic charges (and smaller ions) give stronger attraction and a higher melting point.

Electrical conductivity. A current needs **charged particles that can move**.

Structure	Solid	Liquid (molten)	Why
Giant metallic	conducts	conducts	delocalised electrons move through the structure
Giant ionic	does not	conducts	ions are fixed in the solid; free to move when molten or dissolved
Giant covalent	does not (graphite does)	does not	all electrons in bonds; graphite has delocalised electrons along its layers
Simple molecular	does not	does not	molecules have no overall charge; no free electrons or ions

Solubility.

- Many ionic compounds dissolve in water, because the polar water molecules are attracted to the ions and pull them out of the lattice. Some do not: magnesium oxide and aluminium oxide are insoluble.
- Simple molecular substances that can form hydrogen bonds with water (small alcohols, ammonia) dissolve in water. Non-polar ones (iodine, C₆₀) hardly dissolve in water but dissolve in non-polar solvents such as hexane.
- Giant covalent substances and metals do not dissolve in water or in non-polar solvents.

Other properties. Metals are **malleable** (can be hammered into shape): layers of ions can slide over each other and the delocalised electrons still hold them together. Ionic solids are **brittle**: a small shift brings ions of the same charge side by side, and they repel. Diamond is very hard because every atom is locked in by four strong bonds.

Key facts and formulas

"Given" means it is in the 9701 Data Booklet. "Learn it" means you must remember it.

Formula	Status
$pV = nRT$	Learn it
$R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$	Given
Molar volume of a gas: $22.4 \text{ dm}^3 \text{ mol}^{-1}$ at s.t.p., $24.0 \text{ dm}^3 \text{ mol}^{-1}$ at room conditions	Given
$T (\text{K}) = \theta (^\circ\text{C}) + 273$	Learn it
$1 \text{ dm}^3 = 10^{-3} \text{ m}^3$; $1 \text{ cm}^3 = 10^{-6} \text{ m}^3$; $1 \text{ kPa} = 1000 \text{ Pa}$	Learn it
$n = \frac{m}{M}$	Learn it
$M_r = \frac{mRT}{pV}$	Learn it
$M_r = \frac{\rho RT}{p} (\rho \text{ in g m}^{-3})$	Learn it
At fixed V and T : $p \propto n$; at fixed n and T : $p_1 V_1 = p_2 V_2$	Learn it
Ideal gas: particles have zero volume and no intermolecular forces	Learn it
Most ideal: high temperature, low pressure, weak intermolecular forces	Learn it

How to do it

In the Paper 1 and Paper 2 questions we have tagged for this topic (70 questions, 2021–2025), the types came up this often. 10 of the 70 questions mix more than one type, so the counts add up to 81.

Question type	Questions
Use $pV = nRT$ to find p , V , n , mass or density	13
Find M_r or identify a gas from gas data	12
Ideal gases: assumptions, pressure and real-gas behaviour	12
Explain physical properties from structure and bonding	11
Describe or recognise a lattice structure	10
Deduce structure and bonding from given properties	9
Gas volumes from reacting amounts	7
Changing conditions: connected vessels, mixtures and graphs	7

1. Use $pV = nRT$ to find p , V , n , mass or density (13 questions)

1. **List the data and convert every quantity to SI units** before you touch the equation: Pa, m^3 , K, mol.
2. **Rearrange first**, then substitute: $n = \frac{pV}{RT}$, $V = \frac{nRT}{p}$, $p = \frac{nRT}{V}$.
3. **Calculate** and convert the answer back to the unit asked for (m^3 to dm^3 : $\times 1000$; m^3 to cm^3 : $\times 10^6$).

Variations you will meet:

- **Mass of gas:** find n , then $m = n \times M$.
- **Density:** take 1 mol (or any amount), find its volume and divide mass by volume; or use $\rho = \frac{pM}{RT}$ (in g m^{-3} , then divide by 1000 for g dm^{-3}).
- **Multiple choice with expressions:** the options differ in where the powers of ten and 273 go. Write the correct expression yourself first, with every unit converted, and then match it.
- **Concentration of a dissolved gas:** find n from its gas volume, then divide by the solution volume in dm^3 .

2. Find M_r or identify a gas from gas data (12 questions)

1. Convert to SI units.
2. Find $n = \frac{pV}{RT}$.
3. $M_r = \frac{\text{mass}}{n}$; or in one step $M_r = \frac{mRT}{pV}$.
4. **Identify:** compare your M_r with the options' M_r values (use the Periodic Table's A_r values).

Variations you will meet:

- **From density:** $M_r = \frac{\rho RT}{p}$, remembering ρ in g dm^{-3} must become g m^{-3} .
- **Molecular formula of an element's vapour** (for example sulfur as S_2 , S_4 or S_8): divide M_r by A_r to get the number of atoms per molecule.
- **Hydrocarbon or chloride from M_r :** match M_r to a formula, for example which Period 3 chloride or which hydrocarbon (then its skeletal formula).
- **Volatile liquid in a gas syringe:** the volume *increase* of the syringe is the volume of the vapour.

3. Ideal gases: assumptions, pressure and real-gas behaviour (12 questions)

- **"State two assumptions of an ideal gas":** particles have **zero (negligible) volume**; there are **no intermolecular forces**. (Not "no mass".)
- **"Conditions for ideal behaviour":** **high temperature and low pressure**. Say both.
- **"Which gas is most ideal?"** The one with the weakest intermolecular forces: a small non-polar molecule or a noble gas with few electrons (helium best). **Least ideal:** hydrogen-bonded (steam, ammonia, HF) or polar, or with many electrons.
- **"Explain why a gas is not ideal at very high pressure":** the molecules are close together, so their own volume is a significant part of the total volume, and intermolecular forces become significant. (A strong bond *inside* a molecule has nothing to do with it.)
- **"Explain why gas X is more ideal than gas Y":** compare their intermolecular forces. For example, N_2 (non-polar, only instantaneous dipole–induced dipole forces) is more ideal than CO (polar, which adds permanent dipole–permanent dipole forces), even though they have the same number of electrons.
- **Origin of pressure:** molecules collide with the walls of the container; pressure is the force of these collisions per unit area.

- **Particles in solids, liquids and gases:** solid, particles vibrate about fixed positions, fixed shape; liquid, particles close together but moving past each other, hard to compress; gas, particles far apart and moving fast in random directions, easy to compress.

4. Explain physical properties from structure and bonding (11 questions)

1. **Name the structure and bonding**, including "giant" or "simple".
2. **Name what must be overcome** (for melting or boiling) or **what carries the charge** (for conductivity).
3. **Link to strength or mobility:** "many strong covalent bonds, needing a lot of energy to break" or "weak intermolecular forces, needing little energy".

Variations you will meet:

- **Comparing melting points of two substances:** say what is broken in **each**, and which needs more energy. Metal against ionic compound: compare metallic bonds with ionic bonds. Phosphorus (P_4) against sulfur (S_8): both simple molecular; S_8 has more electrons, so stronger instantaneous dipole–induced dipole forces.
- **Allotropes** (white phosphorus P_4 against red phosphorus; C_{60} against graphite): simple molecular melts low because only intermolecular forces are overcome; giant covalent melts high because covalent bonds must break.
- **"In which substance are covalent bonds broken on melting?":** only a giant covalent one, such as SiO_2 .
- **Identify the strongest force broken on melting** for each substance in a table: giant ionic, ionic bond; giant covalent, covalent bond; simple molecular, the strongest intermolecular force (hydrogen bond, permanent dipole–permanent dipole or instantaneous dipole–induced dipole).
- **Evaluate a hypothesis** (for example "stronger covalent bonds give higher melting points"): check whether the data really compare the same thing. If most of the substances break intermolecular forces, not covalent bonds, on melting, the data cannot test it.
- **Conductivity of graphite and diamond:** graphite has delocalised electrons that move along the layers; diamond has none, because all four outer electrons are used in bonds.
- **Volatility of liquids:** the liquid with stronger intermolecular forces (for example more hydrogen bonds per molecule, as in water) needs more energy for molecules to escape, so it evaporates less readily and boils higher.

5. Describe or recognise a lattice structure (10 questions)

- **Name the type:** giant ionic, simple molecular, giant molecular (giant covalent), giant metallic. Always include "giant" or "simple".
- **Count neighbours:** $NaCl$ is 6 : 6; in SiO_2 each Si has 4 O and each O has 2 Si, all covalent; in diamond each C has 4; in graphite each C has 3.
- **Pick the right diagram:** in the sodium chloride lattice the particles are **ions**, not atoms, and ions alternate in all three directions.
- **Count the simple molecular ones in a list:** iodine and C_{60} are simple molecular; graphite and diamond are giant.

- **Draw metallic bonding:** a regular arrangement of positive ions, labelled, with delocalised electrons, labelled, between them.
- **A new layered solid** (such as boron nitride): if it has hexagonal layers held by weak forces, the similar solid is graphite.

6. Deduce structure and bonding from given properties (9 questions)

Use the properties in this order:

1. **Conducts as a solid** (and as a liquid)? Giant metallic (or graphite).
2. **Conducts only when molten or dissolved?** Giant ionic.
3. **Never conducts, high melting point** (usually over 1000 °C)? Giant covalent.
4. **Never conducts, low melting point?** Simple molecular.

Then use **solubility** to choose between two options of the same type, for example an insoluble ionic oxide (Al₂O₃, MgO) against a soluble ionic salt (NaF, NaCl).

Variations you will meet:

- **A change of form**, such as a brittle non-conductor turning into a malleable conductor: giant covalent to giant metallic.
- **A table to complete** for elements or compounds: state, bonding (covalent, ionic, metallic) and structure (giant or simple).
- **Chlorides and oxides of Period 3:** NaCl and MgO giant ionic; SiO₂ giant covalent; SiCl₄, PCl₃, SO₂ simple molecular.
- **A solid that melts on gentle heating and reacts with water:** simple molecular with covalent bonding.

7. Gas volumes from reacting amounts (7 questions)

1. **Moles of the thing you know** ($n = \frac{m}{M}$, or $n = c \times V$ for a solution). If both reactants are given, find the limiting one.
2. **Use the equation's ratio** to get the moles of gas. Count **every gas** made if the question says "total volume of gas", and check the temperature: water is a gas above 100 °C.
3. **Use $pV = nRT$** to turn moles of gas into a volume (or the other way round).

Variations you will meet:

- **Working backwards** from a gas volume to the moles or volume of a reactant (for example the volume of acid used).
- **Change in pressure in a sealed vessel:** at fixed V and T , pressure is proportional to total moles of gas. Let x mol react, write the total moles after in terms of x , and set the ratio equal to the pressure ratio.
- **Percentage of a gas used up:** moles of gas at the start from $pV = nRT$; moles used from the equation.

8. Changing conditions: connected vessels, mixtures and graphs (7 questions)

1. **Moles are conserved** when vessels are connected. Find the moles in each vessel (or use pV for each when the temperature is the same).
2. **The final volume is the total volume** of all the connected vessels (an evacuated one included).
3. **Use the final temperature** with the total moles and total volume: $p = \frac{n_{\text{total}}RT}{V_{\text{total}}}$.

At one fixed temperature there is a quick form: $p_{\text{final}} = \frac{p_1V_1 + p_2V_2}{V_1 + V_2}$. If the temperature then changes, multiply by $\frac{T_2}{T_1}$.

Variations you will meet:

- **Masses given instead of pressures:** use the masses to get moles; the starting pressures may then not be needed.
- **Mixtures of gases:** each gas's partial pressure is (its moles $\div$ total moles) $\times$ total pressure.
- **Graphs:** at constant T , p against V is a curve (a hyperbola); p against $\frac{1}{V}$ is a straight line through the origin; at constant n and V , p against T in kelvin is a straight line through the origin.
- **A reaction removes or makes gas** when a tap is opened: think about how the moles of gas change.

Worked examples

The bold codes show where marks are earned, as in Cambridge mark schemes. The number is how many marks.

- **M1** method mark: for a correct method, even if the numbers slip.
- **A1** accuracy mark: for a correct answer; it needs the M mark before it.
- **B1** mark for a correct result on its own, with no method needed.

Example 1

A volatile liquid, Z , is injected into a gas syringe heated to 90°C . All of it evaporates. The mass of Z used is 0.391 g and the volume of vapour produced is 162 cm^3 . The pressure is 98.5 kPa .

- (a) Calculate the relative molecular mass of Z . Show your working. **[3]**
- (b) Z is one of propan-1-ol ($M_r = 60.0$), butan-1-ol ($M_r = 74.0$) and pentanoic acid ($M_r = 102.0$). Identify Z . **[1]**
- (a) Convert: $T = 90 + 273 = 363\text{ K}$, $V = 162 \times 10^{-6}\text{ m}^3$, $p = 98\,500\text{ Pa}$. **B1** for all three conversions.

$$n = \frac{pV}{RT} = \frac{98\,500 \times 162 \times 10^{-6}}{8.31 \times 363} = 0.00529\text{ mol}$$

M1 for the correct rearrangement and substitution.

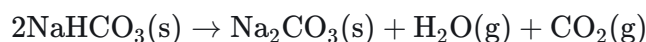
$$M_r = \frac{0.391}{0.00529 \dots} = 73.9$$

A1. (Keep the unrounded n in your calculator; rounding it early changes the last figure.)

(b) Butan-1-ol, because its M_r of 74.0 is closest to 73.9. **B1**

Example 2

Sodium hydrogencarbonate decomposes when heated.



4.20 g of NaHCO_3 ($M_r = 84.0$) is heated until it has all decomposed.

(a) Calculate the total volume of gas, in cm^3 , produced at 150°C and 100 kPa. **[3]**

(b) The gas is cooled to room conditions. Calculate the volume of gas left, in dm^3 , using the molar volume of a gas at room conditions. **[1]**

(a) $n(\text{NaHCO}_3) = \frac{4.20}{84.0} = 0.0500$ mol, so 0.0250 mol of H_2O and 0.0250 mol of CO_2 : 0.0500 mol of gas altogether. **M1** for the 2 : 1 : 1 ratio and counting both gases.

$$V = \frac{nRT}{p} = \frac{0.0500 \times 8.31 \times 423}{100\,000} = 1.76 \times 10^{-3} \text{ m}^3$$

M1 for $pV = nRT$ with T in kelvin.

$V = 1760 \text{ cm}^3$ (3 s.f.). **A1**

(b) At room temperature the water condenses, so only the 0.0250 mol of CO_2 is still a gas: $0.0250 \times 24.0 = 0.600 \text{ dm}^3$. **B1**

Example 3

Vessel A has a volume of 1.50 dm^3 and holds oxygen at 240 kPa. Vessel B has a volume of 4.50 dm^3 and holds nitrogen at 80 kPa. Both are at 300 K and they are joined by a closed tap. Assume both gases behave ideally.

(a) The tap is opened and the temperature is kept at 300 K. Calculate the final pressure. **[2]**

(b) The whole apparatus is then heated to 400 K. Calculate the new total pressure and the partial pressure of oxygen. **[3]**

(a) The moles are conserved and the temperature does not change, so pV adds up:

$$p = \frac{(240 \times 1.50) + (80 \times 4.50)}{1.50 + 4.50} = \frac{360 + 360}{6.00} = 120 \text{ kPa}$$

M1 for adding the pV values and dividing by the **total** volume. **A1**

(b) At fixed n and V , $p \propto T$: $p = 120 \times \frac{400}{300} = 160 \text{ kPa}$. **M1 A1**

(Check with moles: $n(\text{O}_2) = \frac{240\,000 \times 1.50 \times 10^{-3}}{8.31 \times 300} = 0.1444 \text{ mol}$, the same as $n(\text{N}_2)$, so 0.289 mol in $6.00 \times 10^{-3} \text{ m}^3$ at 400 K gives 160 kPa .)

The two gases have equal moles, so oxygen's partial pressure is half the total: 80 kPa . **B1**

Example 4

The table shows some properties of three solids, X, Y and Z.

Solid	Melting point / °C	Conducts as a solid?	Conducts when molten?	Solubility in water
X	1085	yes	yes	insoluble
Y	114	no	no	very slightly soluble; dissolves in hexane
Z	771	no	yes	soluble

(a) Deduce the type of structure in each of X, Y and Z. **[3]**

(b) Explain why Z conducts electricity when molten but not when solid. **[2]**

(c) Explain, in terms of structure and bonding, why Y has a much lower melting point than Z. **[3]**

(a) X: giant metallic (conducts as a solid and a liquid). **B1** Y: simple molecular (low melting point, never conducts). **B1** Z: giant ionic (conducts only when molten, high melting point, soluble). **B1**

(b) In the solid the ions are held in fixed positions in the lattice and cannot move. **B1** When molten the ions are free to move and carry the charge. **B1**

(c) Melting Y only overcomes the weak intermolecular forces between its molecules; its covalent bonds do not break. **B1** Melting Z means overcoming the strong electrostatic attraction between oppositely charged ions throughout the giant lattice. **B1** Much more energy is needed for Z, so its melting point is higher. **B1**

Common mistakes

- **Mixing units.** The most common error in examiner reports every year: volumes left in cm^3 or dm^3 , pressure in kPa , or temperature in $^\circ\text{C}$. Convert everything to Pa , m^3 and K before substituting.
- **Rearranging wrongly.** Write the rearranged equation in symbols (for example $n = \frac{pV}{RT}$) before putting in numbers.
- **Rounding too early.** Rounding n to one or two figures before dividing a mass by it gives a noticeably different M_r . Keep full calculator values until the end.
- **Stopping at n .** After finding the moles of gas, examiners see candidates forget to use the reaction's mole ratio to get back to the reactant (or to include every gas made).
- **"Ideal gas particles have no mass."** Wrong: they have mass but **no volume**. The other assumption is **no intermolecular forces**.

- **Getting the conditions the wrong way round.** Most ideal at **high** temperature and **low** pressure. When asked for conditions, give temperature and pressure, not the assumptions.
- **Blaming a strong bond for non-ideal behaviour.** A triple bond inside N_2 does not matter; it is the molecules' own volume and the attractions **between** molecules.
- **Naming only the strongest intermolecular force** when asked for all of them. A polar molecule such as CO has instantaneous dipole–induced dipole forces **and** permanent dipole–permanent dipole forces.
- **Talking about intermolecular forces in a giant covalent structure.** Melting diamond or SiO_2 breaks **covalent bonds**. (Graphite's layers are the one place in these structures where weak forces matter, and that is about softness, not melting.)
- **Leaving out "giant".** "Metallic lattice" or "ionic lattice" is not enough: write **giant** metallic, **giant** ionic.
- **Metallic diagrams without labels**, or with the ions called "atoms", "nuclei" or "protons". Label **positive ions** and **delocalised electrons**.
- **Saying delocalised electrons exist but not that they move.** Conduction needs the electrons (or ions) to **move through the structure**.
- **Not using the information given.** If the question says a substance is like copper, or gives a table, your answer must use it.

How to get the marks

- **Calculations: show your working.** A method mark usually goes on converting to K and m^3 (sometimes one mark each), and another on the correct rearrangement. A final answer with no working may score nothing if it is wrong.
- **Give the unit and the significant figures asked for.** "Give your answer to 3 significant figures" costs a mark if ignored. A volume "in m^3 " must be in m^3 .
- **Use the Data Booklet values:** $R = 8.31$, and 22.4 or $24.0 \text{ dm}^3 \text{ mol}^{-1}$ only when the question sets s.t.p. or room conditions.
- **Error carried forward:** if you use a wrong answer from an earlier part correctly, you still earn the later marks. If you can't do a part, use the value the question offers instead.
- **"Explain" for melting points** needs three things: the structure, what is broken or overcome, and the energy needed ("many strong ... bonds, a lot of energy"). Compare **both** substances.
- **"Explain" for conductivity** needs the charge carrier (delocalised electrons or ions) **and** that it is free to move.
- **"State" or "name"** needs only the answer, but use the exact terms: *giant molecular*, *simple molecular*, *instantaneous dipole–induced dipole*, *delocalised electrons*.
- **"Suggest"** means apply what you know to something unfamiliar (a new allotrope, a new chloride); give a reason from the data.
- **Multiple-choice gas questions:** work the answer out yourself, then find it. The wrong options are built from the usual slips ($^{\circ}\text{C}$ for K, a wrong power of ten, a missed mole ratio).

Quick check

1. Calculate the volume, in cm^3 , occupied by 0.640 g of oxygen gas, O_2 , at 27°C and 150 kPa.
2. A gas has a density of 1.75 g dm^{-3} at 60°C and 110 kPa. Calculate its M_r .
3. 0.540 g of aluminium reacts with an excess of hydrochloric acid: $2\text{Al} + 6\text{HCl} \rightarrow 2\text{AlCl}_3 + 3\text{H}_2$. Calculate the volume of hydrogen, in cm^3 , at 30°C and 100 kPa.
4. Explain why hydrogen fluoride behaves much less like an ideal gas than argon under the same conditions.
5. Describe the structure of solid copper, and use it to explain why copper conducts electricity and is malleable.

Answers

1. $n = \frac{0.640}{32.0} = 0.0200 \text{ mol}$. $V = \frac{0.0200 \times 8.31 \times 300}{150\,000} = 3.32 \times 10^{-4} \text{ m}^3 = 332 \text{ cm}^3$.
2. $\rho = 1750 \text{ g m}^{-3}$ and $T = 333 \text{ K}$. $M_r = \frac{\rho RT}{p} = \frac{1750 \times 8.31 \times 333}{110\,000} = 44.0$.
3. $n(\text{Al}) = \frac{0.540}{27.0} = 0.0200 \text{ mol}$, so $n(\text{H}_2) = 0.0200 \times \frac{3}{2} = 0.0300 \text{ mol}$. $V = \frac{0.0300 \times 8.31 \times 303}{100\,000} = 7.55 \times 10^{-4} \text{ m}^3 = 755 \text{ cm}^3$.
4. An ideal gas has no intermolecular forces. HF molecules are polar and form hydrogen bonds with each other, so there are strong attractions between them. Argon atoms have only weak instantaneous dipole–induced dipole forces, so argon is much closer to ideal.
5. Copper is a giant metallic lattice: a regular arrangement of positive copper ions in a sea of delocalised electrons. It conducts because the delocalised electrons can move through the structure and carry charge. It is malleable because layers of ions can slide over each other while the delocalised electrons keep holding them together, so the metallic bonding is not broken.

Past questions to try

- **9701/11/M/J/25 Q13**: connected containers, with masses given and a new temperature.
- **9701/22/M/J/25 Q1**: graphite and diamond, and the force broken when each oxide melts.
- **9701/22/M/J/21 Q2**: a giant ionic melting point, ideal gas assumptions and why nitrogen is not ideal at high pressure.
- **9701/22/F/M/21 Q1**: ideal gas assumptions, then the pressure of krypton from $pV = nRT$.