



iBLaurel

Structure 1.5.3 & 1.5.4 Molar Volume and the Ideal Gas Equation

The gas laws, $PV = nRT$ and the combined gas law

IB Diploma Programme Chemistry · First assessment 2025
Notes · Formula sheet · Graph work · Exam technique · 16 solved examples · 20 MCQs · Paper 1B & Paper 2 practice · Full mark schemes

SL + HL core content

HL Stretch harder application

Forward link 1.4 / 2.2 / Tools

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How to use this booklet

READ FIRST

This booklet covers the final two statements of **Structure 1.5 — Ideal gases** of the IB Diploma Programme Chemistry guide (first assessment 2025). It is the companion to the iBLaurel booklet on Structure 1.5.1 and 1.5.2, which covers the assumptions of the model and the deviations of real gases.

SYLLABUS STATEMENTS COVERED

Structure 1.5.3 — The molar volume of an ideal gas is a constant at a specific temperature and pressure.

Assessment skill: Investigate the relationship between temperature, pressure and volume for a fixed mass of an ideal gas and analyse graphs relating these variables.

Guidance: The names of specific gas laws will not be assessed. The value for the molar volume of an ideal gas at STP is given in the data booklet.

Structure 1.5.4 — The relationship between the pressure, volume, temperature and amount of an ideal gas is shown in the ideal gas equation $PV = nRT$ and the combined gas law $P_1V_1/T_1 = P_2V_2/T_2$.

Assessment skill: Solve problems relating to the ideal gas equation.

Guidance: Units of volume and pressure should be SI only. The value of the gas constant R , the ideal gas equation and the combined gas law are given in the data booklet.

What the level labels mean

Label	Meaning
SL + HL	Common core. Structure 1.5 is entirely common core — there is no AHL-only (HL-only) content in 1.5.3 or 1.5.4. Every SL and every HL student is examined on all of it.
HL Stretch	Same content, harder question style. Multi-step calculations, gas stoichiometry combined with limiting reactants, graph gradients and error analysis. SL students aiming for a 7 should work through these too.
Forward / back link	Connections to Structure 1.4 (the mole, Avogadro's law, concentration), Structure 1.5.2 (real-gas deviation) and the Tools and Inquiry strands (graphing, uncertainty, experimental determination of molar mass).

THE ONE-LINE SUMMARY OF THIS WHOLE TOPIC

Two equations do everything. Use **$PV = nRT$** when you are asked about *how much* gas there is (moles, mass, molar mass, density) at *one* set of conditions. Use the **combined gas law** when a *fixed amount* of gas moves from one set of conditions to another and the amount never appears in the question.

Part A — Structure 1.5.3: Molar volume

SL + HL

A1 From Avogadro's law to molar volume SL + HL

Assumption 1 of the ideal gas model says the particles themselves occupy a negligible volume. It follows that the volume a gas occupies depends on the **number** of particles and on the conditions — not on how big or heavy the particles are. That is **Avogadro's law** (Structure 1.4.6):

AVOGADRO'S LAW

Equal volumes of all gases, measured at the same temperature and pressure, contain **equal numbers of molecules**.

Turn that round and you have the statement of Structure 1.5.3: if equal numbers of particles occupy equal volumes, then *one mole* of any gas must occupy the *same* volume at a given temperature and pressure. That volume is the **molar volume**.

DEFINITION

The **molar volume** of a gas, V_m , is the volume occupied by one mole of the gas at a stated temperature and pressure. For an ideal gas it is the same for every gas at those conditions.

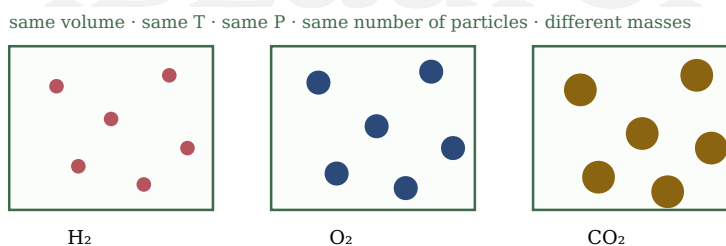


Figure 1 Avogadro's law. The three containers hold equal volumes at the same temperature and pressure, so they contain equal numbers of particles — even though the particles differ in size and mass. This is why one molar volume serves every gas.

A2 The value you must be able to use SL + HL

$$V_m = 2.27 \times 10^{-2} \text{ m}^3 \text{ mol}^{-1} = 22.7 \text{ dm}^3 \text{ mol}^{-1} \text{ at STP}$$

STP = 273 K (0 °C) and 100 kPa (10^5 Pa). Given in Section 2 of the data booklet.

COMMON ERROR — 22.4 OR 22.7?

22.7 dm³ mol⁻¹ is the value for the current syllabus. The older figure of 22.4 dm³ mol⁻¹ corresponds to a pressure of 101.3 kPa (1 atm), which the IB no longer uses as standard pressure. Older textbooks, websites and worksheets still quote 22.4 — using it will cost you accuracy marks. Check the data booklet if in doubt.

Note also that **STP is not the same as "standard state"** used in thermodynamics (Reactivity 1.1), which specifies 298 K. Two different standards, two different temperatures.

A3 The three calculations molar volume gives you SL + HL

$$n = V / V_m \quad V = n \times V_m \quad \rho = M / V_m$$

Volume and molar volume must be in the **same** unit — dm³ with 22.7, or m³ with 2.27×10^{-2} .

You are given	You want	Route
Volume of gas at STP	Amount in mol	Divide by 22.7 dm ³ mol ⁻¹ . This replaces $n = m/M$ — you do not need the molar mass at all.
Mass of gas	Volume at STP	$n = m/M$, then $V = n \times 22.7 \text{ dm}^3$.
Density at STP (g dm ⁻³)	Molar mass	$M = \rho \times 22.7$. The density of one mole occupying 22.7 dm ³ is $M/22.7$, so the relation reverses directly.
Moles from an equation	Volume of gas produced	Standard mole-ratio stoichiometry, then $\times 22.7 \text{ dm}^3 \text{ mol}^{-1}$ at the end. HL Stretch

EXAM TIP — MOLAR VOLUME ONLY WORKS AT THE STATED CONDITIONS

22.7 dm³ mol⁻¹ is valid **only at 273 K and 100 kPa**. If the question gives any other temperature or pressure, you have two choices: (i) use $PV = nRT$ directly (usually quicker), or (ii) convert the molar volume to the new conditions with the combined gas law. Never apply 22.7 to a question set at 25 °C.

A4 Molar volume at other conditions SL + HL

Because $V_m = RT/P$, the molar volume rises with temperature and falls with pressure. Useful values:

Conditions	Molar volume	Comment
273 K, 100 kPa (STP)	22.7 dm³ mol⁻¹	Data booklet value. Learn this one.

298 K, 100 kPa (room conditions)	$24.8 \text{ dm}^3 \text{ mol}^{-1}$	Obtained from $22.7 \times 298/273$ or from RT/P . Not given to you — derive it if needed.
273 K, 101.3 kPa (1 atm)	$22.4 \text{ dm}^3 \text{ mol}^{-1}$	The old value. Shown here only so you recognise where it comes from.

SENSE CHECK

A basketball fits loosely inside a box of 22.7 dm^3 . If your answer for the volume of a few moles of gas comes out in cubic centimetres, or in hundreds of cubic metres, you have almost certainly slipped a factor of 1000 somewhere.



Part B — Investigating and graphing the gas laws

SL + HL

The assessment skill for Structure 1.5.3 is to *investigate* the relationships between P , V and T for a fixed mass of gas and to *analyse graphs* relating them. The specific gas laws are **not named** in examinations — but the relationships and their graph shapes certainly are.

B1 The three relationships SL + HL

Held constant	Relationship	Explanation from the model (see Structure 1.5.1)
T and n (named: Boyle's law)	$P \propto 1/V$ $PV = \text{constant}$	A larger volume means particles travel further between wall collisions, so collisions are less frequent and the pressure falls. Speed is unchanged because T is unchanged.
P and n (named: Charles' law)	$V \propto T(K)$ $V/T = \text{constant}$	Heating gives each particle more kinetic energy, so each collision is more energetic. To keep the pressure constant the collision frequency must fall, which requires the volume to increase.
V and n (named: Gay-Lussac's law)	$P \propto T(K)$ $P/T = \text{constant}$	Heating at fixed volume makes collisions both more frequent and more energetic, so the pressure rises.

B2 Graph shapes you must recognise instantly SL + HL

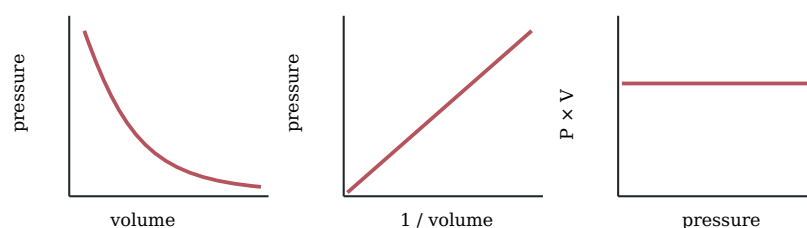


Figure 2 Constant temperature. P against V is a curve (a hyperbola) approaching but never touching the axes; P against $1/V$ is a straight line through the origin; PV against P is horizontal. The last two are the tests for ideal behaviour.

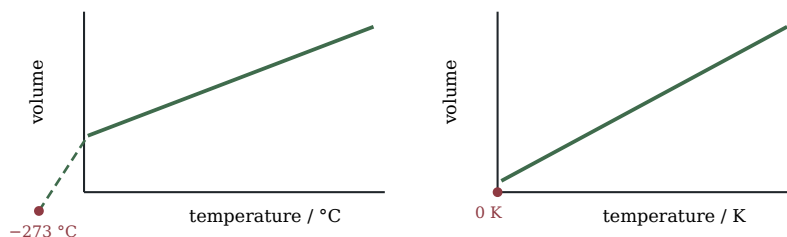


Figure 3 Constant pressure. Against Celsius temperature the line is *linear* with a positive intercept, extrapolating to zero volume at $-273\text{ }^{\circ}\text{C}$. Against absolute temperature it is *proportional* — a straight line through the origin. The same pair of shapes applies to pressure against temperature at constant volume.

Plot	Shape for an ideal gas	What it tells you
P against V	Curve (hyperbola), asymptotic to both axes	Inverse proportionality. Hard to judge by eye — do not use it to test ideality.
P against $1/V$	Straight line through the origin	The proper test of $P \propto 1/V$. Gradient = nRT (constant).
PV against P	Horizontal straight line	Constant PV confirms ideal behaviour at constant T . A sloping line indicates deviation (Structure 1.5.2).
V against $\theta / ^{\circ}\text{C}$	Straight line, positive intercept	Linear but <i>not</i> proportional. Extrapolate to $V = 0$ to find absolute zero.
V against T / K	Straight line through the origin	Direct proportionality. Gradient = nR/P .
P against T / K	Straight line through the origin	Direct proportionality at constant V . Gradient = nR/V .

COMMON ERROR — "LINEAR" AND "PROPORTIONAL" ARE NOT THE SAME WORD

A **linear** relationship gives a straight line. A **proportional** relationship gives a straight line *through the origin*. V against $^{\circ}\text{C}$ is linear only; V against K is proportional. Examiners award the mark for the correct word, so use them precisely.

B3 Determining absolute zero from experimental data

SL + HL

Tools 2/3

A classic practical: measure P (or V) for a fixed mass of gas at several Celsius temperatures, plot the line of best fit, and extrapolate back to the temperature axis.

- Plot the data and draw the best-fit line; obtain its equation, $y = mx + c$.
- Set $y = 0$ and solve: $x = -c/m$. This is the predicted value of absolute zero in $^{\circ}\text{C}$.
- Compare with the accepted value, $-273\text{ }^{\circ}\text{C}$, and calculate the percentage error:

$$\% \text{ error} = \frac{|\text{experimental} - \text{accepted}|}{|\text{accepted}|} \times 100.$$

4. Comment: highly linear data with an offset value indicates a **systematic** error (gauge not zeroed, a leak, gas in connecting tubing not at bath temperature). Scattered data indicates **random** error.

EXAM TIP — THE EXTRAPOLATION IS THE POINT

The dashed part of the line in Figure 3 is an **extrapolation**: data at temperatures approaching 0 K cannot be obtained, because every real gas condenses long before then. Saying so is often worth a mark in "comment on" or "evaluate" questions.

B4 Gradients and intercepts you might be asked for HL Stretch

Plot	Gradient equals	Use
P against $1/V$ (constant T, n)	nRT	Read the gradient, then find n if T is known — or find T if n is known.
V against $T(K)$ (constant P, n)	nR/P	Determine the amount of gas from the gradient.
P against $T(K)$ (constant V, n)	nR/V	Determine n , or the volume of the vessel.
P against $\theta(^{\circ}C)$	nR/V (unchanged)	Intercept = the pressure at $0^{\circ}C$. Changing from K to $^{\circ}C$ shifts the line sideways but does not alter its gradient.

Part C — Structure 1.5.4: $PV = nRT$

SL + HL

C1 The combined gas law SL + HL

The three relationships of Part B apply to the same fixed mass of gas, so they can be merged into one statement:

$$PV/T = \text{a constant} \Rightarrow P_1V_1/T_1 = P_2V_2/T_2$$

Subscript 1 = initial conditions, subscript 2 = final conditions. Given in Section 1 of the data booklet.

- The amount of gas must be **fixed** — no gas added, removed or lost.
- T must be in **kelvin**. P and V may be in any units, provided the **same** unit is used on both sides. This is the one place where you are allowed to keep kPa and dm^3 without converting, because the units cancel.
- If one variable is held constant, cancel it from both sides. Constant temperature reduces the expression to $P_1V_1 = P_2V_2$; constant pressure gives $V_1/T_1 = V_2/T_2$; constant volume gives $P_1/T_1 = P_2/T_2$.

C2 The ideal gas equation SL + HL

Combining $PV/T = \text{constant}$ with Avogadro's law (the constant is proportional to the amount of gas, n) gives $PV/T \propto n$. Introducing the proportionality constant R :

$$PV = nRT \quad R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$$

R is the **universal** gas constant — it does not depend on the identity of the gas, because the ideal model contains nothing about chemical identity.

You can confirm the value of R from the molar volume at STP:

$$R = PV/nT = (1.00 \times 10^5 \text{ N m}^{-2} \times 2.27 \times 10^{-2} \text{ m}^3) / (1.00 \text{ mol} \times 273 \text{ K}) = \mathbf{8.31 \text{ J K}^{-1} \text{ mol}^{-1}}$$

UNITS — THE SINGLE BIGGEST SOURCE OF LOST MARKS IN STRUCTURE 1.5

With $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ you **must** use strict SI units:

Quantity	Required	Conversion
Pressure	Pa	kPa $\rightarrow$ Pa: $\times 10^3$. atm $\rightarrow$ Pa: $\times 1.013 \times 10^5$.
Volume	m^3	$\text{dm}^3 \rightarrow \text{m}^3$: $\div 10^3$. $\text{cm}^3 \rightarrow \text{m}^3$: $\div 10^6$.
Temperature	K	$^{\circ}\text{C} \rightarrow \text{K}$: $+ 273$.
Amount	mol	$n = m/M$, with m in grams.

The one legitimate shortcut: **kPa with dm^3** may be substituted directly, because $10^3 \times 10^{-3} = 1$. Nothing else cancels.

C3 Choosing the right equation SL + HL

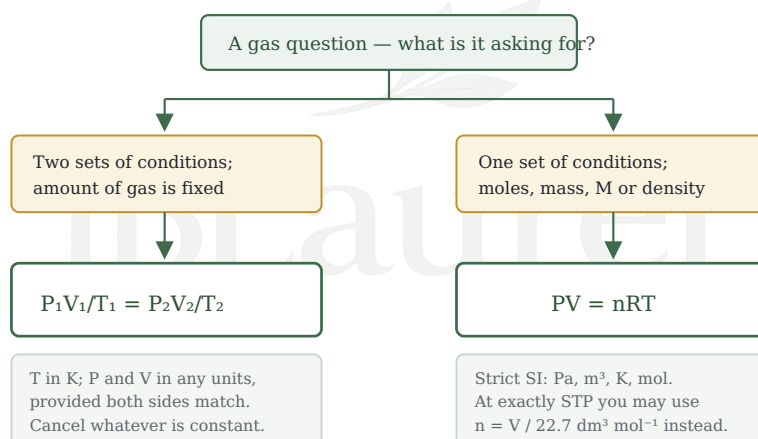


Figure 4 Deciding which relationship to use. The test is whether the amount of gas appears in the question.

C4 Finding molar mass from gas data SL + HL Inquiry 2

Substituting $n = m/M$ into $PV = nRT$ and rearranging gives the standard experimental route to a molar mass:

$$M = mRT / PV \quad \text{and, using density,} \quad M = \rho RT / P$$

In the density form, ρ must be in kg m^{-3} for M in kg mol^{-1} , or in g m^{-3} for M in g mol^{-1} . A density in g dm^{-3} must be multiplied by 1000 first.

Typical experiment: a gas syringe or a weighed flask of known volume is filled with the gas; the mass, volume, temperature and pressure are measured; M follows directly. Points examiners look for when evaluating such an experiment:

- The gas must be **dry** — water vapour adds mass and volume and lowers the apparent M .

- The flask contains **air** before filling; failing to evacuate or to correct for it is a systematic error.
- The value obtained is a little **too high** if the gas has strong intermolecular attractions, because the real volume is smaller than the ideal prediction (Structure 1.5.2).
- The largest percentage uncertainty usually lies in the **mass** of gas, since gases are light — quote uncertainties and identify the dominant one.

C5 Gas stoichiometry

HL Stretch

1.4

Gas questions in Paper 2 are usually stoichiometry questions with a gas step at the end. The reliable sequence is:

1. Write and balance the equation.
2. Convert the given quantity to **moles** (via m/M , $c \times V$, or V/V_m).
3. If two reactants are given, identify the **limiting reactant** by dividing each amount by its coefficient.
4. Use the **mole ratio** to find the moles of the gas required.
5. Convert to volume: $\times 22.7 \text{ dm}^3 \text{ mol}^{-1}$ at STP, or $V = nRT/P$ at any other conditions.

SHORTCUT WORTH KNOWING

When *all* the species involved are gases at the *same* temperature and pressure, the mole ratio *is* the volume ratio, so you can work directly in cm^3 without ever converting to moles. In $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$, 20 cm^3 of nitrogen reacts with 60 cm^3 of hydrogen to give 40 cm^3 of ammonia. This saves several minutes in gas-syringe questions.

C6 How real-gas behaviour affects these calculations

Structure 1.5.2

Everything in this section assumes ideal behaviour. In practice the results are accurate to a few per cent at ordinary conditions, but at high pressure or low temperature the calculated volume or molar mass will be in error. If a question asks you to *evaluate* a result, a sentence on this — naming the failing assumption and the direction of the error — is often the final mark.

Part D — Formula and data sheet

ONE-PAGE REVISION

D1 Given to you in the data booklet

Item	Value / form	Section
Ideal gas equation	$PV = nRT$	1
Combined gas law	$P_1V_1/T_1 = P_2V_2/T_2$	1
Gas constant, R	$8.31 \text{ J K}^{-1} \text{ mol}^{-1}$	2
Molar volume at STP	$2.27 \times 10^{-2} \text{ m}^3 \text{ mol}^{-1} = 22.7 \text{ dm}^3 \text{ mol}^{-1}$	2
Avogadro's constant	$6.02 \times 10^{23} \text{ mol}^{-1}$	2
Relative atomic masses	periodic table	—

D2 Relationships you must supply yourself

Relationship	When it is used
$n = V / V_m$	Amount of gas from volume, at STP only (273 K, 100 kPa).
$n = m / M$	Amount from mass. The bridge between gas data and stoichiometry.
$M = mRT / PV$	Molar mass from a measured mass and volume of gas.
$M = \rho RT / P$	Molar mass from a gas density (ρ in g m^{-3}).
$M = \rho \times 22.7$	Molar mass from a density in g dm^{-3} , at STP only .
$\rho = m / V = PM / RT$	Density of a gas at any conditions. Shows $\rho \propto M$ at fixed T and P .
$T(\text{K}) = \theta(^{\circ}\text{C}) + 273$	Always, before any substitution.
$1 \text{ dm}^3 = 10^{-3} \text{ m}^3$; $1 \text{ cm}^3 = 10^{-6} \text{ m}^3$; $1 \text{ kPa} = 10^3 \text{ Pa}$	Unit conversions. Write them out — do not do them in your head.
STP = 273 K and 100 kPa	Distinguish from "standard state" (298 K) in Reactivity 1.1.

D3 Rearrangements worth memorising

Solve for	Expression	Typical question
n	PV/RT	How many moles of gas are in this container?

V	nRT/P	What volume of gas is produced by this reaction?
P	nRT/V	What is the pressure inside this cylinder?
T	PV/nR	At what temperature would the gas reach this pressure?
M	mRT/PV	Identify the unknown gas.



Part E — Exam technique

SL + HL

E1 A method that works for every calculation in this topic

Step	What to write down
1	List the data with units down the left of your working: $P = \dots$, $V = \dots$, $T = \dots$, $n = \dots$, and mark the unknown with a question mark. Half the marks in a 4-mark calculation are for correct setting-out.
2	Convert everything to SI immediately — before you decide which equation to use. Do not carry a mixture of cm^3 and kPa into the algebra.
3	Choose the equation using Figure 4: two sets of conditions and fixed amount $\rightarrow$ combined gas law; one set of conditions and moles or mass involved $\rightarrow PV = nRT$.
4	Rearrange symbolically first , then substitute. Rearranging with numbers already in place is where sign and inversion errors happen.
5	Check the direction of the answer against physical sense: heating at fixed volume must raise the pressure; compressing must reduce the volume.
6	Round at the end only , to the significant figures of the least precise datum (usually 3), and state the unit . An unlabelled number rarely gets the final mark.

E2 Twelve habits that turn a 5 into a 7

#	Habit
1	Kelvin, every time. Write $T = \theta + 273$ as your first line even when the temperature is already in kelvin, so the habit never fails you.
2	Use 22.7, not 22.4 — and only at 273 K and 100 kPa. At any other conditions, reach for $PV = nRT$.
3	Divide by 10^6 for $\text{cm}^3 \rightarrow \text{m}^3$. The most frequent single arithmetic slip in this topic is dividing by 10^3 instead.
4	Do not mix the two equations. If the amount of gas is constant and unknown, the combined gas law removes the need to find it; introducing R just adds work and risk.
5	Identify the limiting reactant whenever two reactant quantities are given — divide each amount by its coefficient and take the smaller.
6	Use volume ratios directly when every species is a gas at the same T and P . Converting to moles and back wastes time and invites errors.
7	Say "linear" or "proportional" deliberately in graph questions — they are different answers.

8	Quote the gradient's meaning when asked to analyse a graph: nRT for P against $1/V$, nR/P for V against T .
9	Show the substitution , not just the answer. Method marks survive an arithmetic slip; a bare wrong number scores zero.
10	Carry ECF forward . If part (a) went wrong, use your own value consistently in (b) — examiners award error-carried-forward marks.
11	Watch mg and kg . Masses in gas questions are often given in milligrams or kilograms specifically to catch you.
12	Finish evaluation questions with a real-gas sentence — name the failing assumption and the direction of the error.

A 30-SECOND SELF-TEST BEFORE THE EXAM

Can you, from memory: (1) state the molar volume at STP and the conditions it applies to; (2) convert 250 cm^3 to m^3 ; (3) rearrange $PV = nRT$ for molar mass; (4) sketch P against $1/V$ and state its gradient; (5) decide, from a question's wording, which of the two equations to use? If yes, Structure 1.5.3 and 1.5.4 are secure.



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Part F — Solved examples

16 WORKED ANSWERS

Work each question yourself before reading the solution. Values used: $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$, $V_m(\text{STP}) = 22.7 \text{ dm}^3 \text{ mol}^{-1}$, STP = 273 K and 100 kPa.

Example 1 · Volume from mass at STP

SL + HL

Calculate the volume occupied by 0.020 g of helium at STP.

[2]

SOLUTION

$$n = m/M = 0.020 / 4.00 = 5.0 \times 10^{-3} \text{ mol (1)}$$

$$V = n \times V_m = 5.0 \times 10^{-3} \times 22.7 = \mathbf{0.114 \text{ dm}^3} \text{ (114 cm}^3\text{) (1)}$$

EXAMINER'S NOTE

Because the conditions are exactly STP, molar volume is far quicker than $PV = nRT$. Both give the same answer; only one takes ten seconds.

Example 2 · Amount from volume at STP

SL + HL

Calculate the amount, in mol, of oxygen in 500 cm^3 of the gas measured at STP.

[2]

SOLUTION

$$500 \text{ cm}^3 = 0.500 \text{ dm}^3 \text{ (1)}$$

$$n = V/V_m = 0.500 / 22.7 = \mathbf{0.0220 \text{ mol (1)}}$$

EXAMINER'S NOTE

Notice that the identity of the gas was never used. The answer would be identical for any gas — that is the whole content of Structure 1.5.3.

Example 3 · Molar volume at another temperature

SL + HL

The molar volume of a gas at STP is $22.7 \text{ dm}^3 \text{ mol}^{-1}$. Calculate the molar volume at 25.0°C and the same pressure.

[2]

SOLUTION

Pressure is constant, so the combined gas law reduces to $V_1/T_1 = V_2/T_2$.

$$T_1 = 273 \text{ K}, T_2 = 25.0 + 273 = 298 \text{ K (1)}$$

$$V_2 = 22.7 \times 298 / 273 = \mathbf{24.8 \text{ dm}^3 \text{ mol}^{-1} (1)}$$

EXAMINER'S NOTE

The same value comes from $V_m = RT/P = 8.31 \times 298 / 10^5 = 2.48 \times 10^{-2} \text{ m}^3 \text{ mol}^{-1}$. A useful figure to recognise, since room-temperature questions are common — but it is not in the data booklet, so you must derive it.

Example 4 · Density at STP

SL + HL

Calculate the density, in g dm^{-3} , of carbon dioxide at STP.

[2]

SOLUTION

One mole of CO_2 has a mass of 44.01 g and occupies 22.7 dm^3 at STP. (1)

$$\rho = M/V_m = 44.01 / 22.7 = \mathbf{1.94 \text{ g dm}^{-3}} \text{ (1)}$$

EXAMINER'S NOTE

Since $\rho = M/V_m$ and V_m is the same for all gases, at fixed T and P gas density is directly proportional to molar mass. So the denser gas is always the one with the larger M — no calculation needed to compare two gases.

Example 5 · Combined gas law, all three variables changing

SL + HL

A 2.50 dm^3 container of helium at a pressure of 85 kPa was heated from 25°C to 75°C . The container expanded to 2.75 dm^3 . Calculate the final pressure of the helium.

[3]

SOLUTION

$$T_1 = 298 \text{ K}, T_2 = 348 \text{ K}; P_1 = 85 \text{ kPa}, V_1 = 2.50 \text{ dm}^3, V_2 = 2.75 \text{ dm}^3 \text{ (1)}$$

$$P_2 = P_1 V_1 T_2 / (T_1 V_2) = (85 \times 2.50 \times 348) / (298 \times 2.75) \text{ (1)}$$

$$= \mathbf{90.2 \text{ kPa}} \text{ (1)}$$

EXAMINER'S NOTE

No unit conversion of P or V is needed here — kPa and dm^3 appear on both sides and cancel. Only the temperatures must be converted. Sense check: heating raises the pressure, expansion lowers it; heating wins, so the pressure rises slightly. It does.

Example 6 · Solving for an initial temperature

SL + HL

A sample of nitrogen with a volume of 675 cm^3 and a pressure of $1.00 \times 10^5 \text{ Pa}$ was compressed to a volume of 350 cm^3 and a pressure of $2.00 \times 10^5 \text{ Pa}$, at which point its temperature was 27.0°C . Determine its initial temperature.

[3]

SOLUTION

$$T_2 = 27.0 + 273 = 300 \text{ K} \text{ (1)}$$

$$T_1 = T_2 \times P_1 V_1 / (P_2 V_2) = 300 \times (1.00 \times 10^5 \times 675) / (2.00 \times 10^5 \times 350) \text{ (1)}$$

$$= 289 \text{ K} = \mathbf{16.3^\circ\text{C}} \text{ (1)}$$

EXAMINER'S NOTE

Volumes may stay in cm^3 and pressures in Pa because each appears on both sides. Read the final instruction carefully — if the question asks for $^\circ\text{C}$, converting back is worth the mark.

Example 7 · Mass of gas from $PV = nRT$
SL + HL

A helium party balloon has a volume of 18.0 dm^3 . At 25°C the internal pressure is 108 kPa . Calculate the mass of helium in the balloon. [3]

SOLUTION

$$P = 108 \times 10^3 \text{ Pa}, V = 18.0 \times 10^{-3} \text{ m}^3, T = 298 \text{ K} (1)$$

$$n = PV/RT = (108 \times 10^3 \times 18.0 \times 10^{-3}) / (8.31 \times 298) = 1944 / 2476 = 0.785 \text{ mol} (1)$$

$$\text{mass} = nM = 0.785 \times 4.00 = \mathbf{3.14 \text{ g}} (1)$$

EXAMINER'S NOTE

As the conditions are not STP, molar volume cannot be used. Note that substituting 108 and 18.0 directly (kPa with dm^3) gives the same 0.785 mol — the permitted shortcut.

Example 8 · Molar mass from mass and volume
SL + HL

A sample of gas has a volume of 445 cm^3 and a mass of 1.50 g at a pressure of 95.0 kPa and a temperature of 28.0°C . Calculate its molar mass. [3]

SOLUTION

$$P = 95.0 \times 10^3 \text{ Pa}, V = 445 \times 10^{-6} \text{ m}^3, T = 301 \text{ K} (1)$$

$$M = mRT/PV = (1.50 \times 8.31 \times 301) / (95.0 \times 10^3 \times 445 \times 10^{-6}) (1)$$

$$= 3752 / 42.3 = \mathbf{88.8 \text{ g mol}^{-1}} (1)$$

EXAMINER'S NOTE

The commonest error here is dividing the volume by 10^3 rather than 10^6 , which gives $0.0888 \text{ g mol}^{-1}$. A molar mass below that of hydrogen is impossible — always sense-check the magnitude.

Example 9 · Molar mass from density
SL + HL

A gas has a density of 1.65 g dm^{-3} at 27°C and 92.0 kPa . Determine its molar mass and suggest its identity. [3]

SOLUTION

$$1.65 \text{ g dm}^{-3} \text{ means } 1.65 \text{ g occupies } 1.00 \text{ dm}^3 = 1.00 \times 10^{-3} \text{ m}^3. T = 300 \text{ K}. (1)$$

$$M = mRT/PV = (1.65 \times 8.31 \times 300) / (92.0 \times 10^3 \times 1.00 \times 10^{-3}) = 4113 / 92.0 (1)$$

$= 44.7 \text{ g mol}^{-1}$, so the gas is most likely **carbon dioxide** ($M = 44.01 \text{ g mol}^{-1}$). (1)

EXAMINER'S NOTE

The experimental value is about 1.5 % above the true molar mass — consistent with CO_2 deviating slightly from ideal behaviour because of its intermolecular attractions (Structure 1.5.2). Propane, $M = 44.1$, and dinitrogen oxide, $M = 44.0$, are equally acceptable suggestions.

Example 10 · Finding the volume of a container

SL + HL

To find the volume of a flask, it was evacuated so that it contained no gas at all. When 4.40 g of carbon dioxide was introduced, it exerted a pressure of 90.0 kPa at 27 °C. Determine the volume of the flask. [3]

SOLUTION

$$n = 4.40 / 44.01 = 0.100 \text{ mol}; \quad T = 300 \text{ K}; \quad P = 90.0 \times 10^3 \text{ Pa} \quad (1)$$

$$V = nRT/P = (0.100 \times 8.31 \times 300) / (90.0 \times 10^3) = 249.3 / 90\,000 \quad (1)$$

$$= 2.77 \times 10^{-3} \text{ m}^3 = \mathbf{2.77 \text{ dm}^3} \quad (1)$$

EXAMINER'S NOTE

The evacuation step is not decoration: it guarantees that all the measured pressure is due to the CO_2 alone. Questions often ask why it was necessary.

Example 11 · Identifying a gas from its density at STP

SL + HL

An unknown noble gas has a density of 5.84 g dm^{-3} at STP. Calculate its molar mass and identify the gas. [2]

SOLUTION

$$M = \rho \times V_m = 5.84 \times 22.7 = 133 \text{ g mol}^{-1} \quad (1)$$

The noble gas with $M \approx 131.3 \text{ g mol}^{-1}$ is **xenon**. (1)

EXAMINER'S NOTE

At STP this one-line route replaces the whole ideal gas calculation. Recognising that the conditions are exactly STP is worth several minutes across a paper.

Example 12 · Small masses and low pressures

HL Stretch

A 12.1 mg sample of a gas has a volume of 255 cm^3 at a temperature of 25.0 °C and a pressure of 1300 Pa. Determine its molar mass. [3]

SOLUTION

$$m = 12.1 \text{ mg} = 0.0121 \text{ g}; \quad V = 255 \times 10^{-6} \text{ m}^3; \quad T = 298 \text{ K}; \quad P = 1300 \text{ Pa} \quad (1)$$

$$n = PV/RT = (1300 \times 255 \times 10^{-6}) / (8.31 \times 298) = 0.3315 / 2476 = 1.339 \times 10^{-4} \text{ mol (1)}$$

$$M = m/n = 0.0121 / 1.339 \times 10^{-4} = \mathbf{90.4 \text{ g mol}^{-1} \text{ (1)}}$$

EXAMINER'S NOTE

Two traps in one question: milligrams and a pressure already in pascals (so *not* to be multiplied by 10^3). Read the units on the page rather than assuming kPa.

Example 13 · Gas stoichiometry — an airbag

HL Stretch

Sodium azide decomposes as $2\text{NaN}_3(\text{s}) \rightarrow 2\text{Na}(\text{s}) + 3\text{N}_2(\text{g})$.

Determine the volume of nitrogen, in dm^3 , produced by 0.0650 kg of sodium azide at 25.00 °C and 1.08 atm. ($M(\text{NaN}_3) = 65.02 \text{ g mol}^{-1}$; 1 atm = $1.013 \times 10^5 \text{ Pa}$) [4]

SOLUTION

$$n(\text{NaN}_3) = 65.0 / 65.02 = 1.00 \text{ mol (1)}$$

$$\text{From the equation, } n(\text{N}_2) = 1.00 \times 3/2 = 1.50 \text{ mol (1)}$$

$$P = 1.08 \times 1.013 \times 10^5 = 1.094 \times 10^5 \text{ Pa; } T = 298 \text{ K (1)}$$

$$V = nRT/P = (1.50 \times 8.31 \times 298) / (1.094 \times 10^5) = 3.39 \times 10^{-2} \text{ m}^3 = \mathbf{33.9 \text{ dm}^3 \text{ (1)}}$$

EXAMINER'S NOTE

Structure the answer as three separate stages — moles of reactant, mole ratio, gas equation — and label them. Even if the arithmetic slips, the method marks stand. Note the mass in **kilograms**.

Example 14 · Identifying an alkane

SL + HL

An alkane gas has a density of 1.94 g dm^{-3} at STP. Determine its molar mass and identify the alkane. [2]

SOLUTION

$$M = 1.94 \times 22.7 = 44.0 \text{ g mol}^{-1} \text{ (1)}$$

Alkanes have the formula $\text{C}_n\text{H}_{2n+2}$: for $n = 3$, $M(\text{C}_3\text{H}_8) = 44.11 \text{ g mol}^{-1}$. The gas is **propane**. (1)

Example 15 · Volume of gas from a reaction

HL Stretch

Calculate the volume of carbon dioxide, measured at STP, produced when 5.00 g of calcium carbonate decomposes completely: $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$. [3]

SOLUTION

$$M(\text{CaCO}_3) = 40.08 + 12.01 + 3(16.00) = 100.09 \text{ g mol}^{-1} \text{ (1)}$$

$n(\text{CaCO}_3) = 5.00 / 100.09 = 0.0500 \text{ mol}$, and the ratio is 1 : 1, so $n(\text{CO}_2) = 0.0500 \text{ mol}$ (1)

$$V = 0.0500 \times 22.7 = \mathbf{1.13 \text{ dm}^3} \text{ (1)}$$

EXAMINER'S NOTE

Only the CO_2 is a gas — molar volume applies to it alone. Applying $22.7 \text{ dm}^3 \text{ mol}^{-1}$ to a solid or a solution is a serious conceptual error.

Example 16 · Reacting gas volumes without moles

HL Stretch

50 cm^3 of nitrogen is mixed with 200 cm^3 of hydrogen and the mixture is allowed to react completely according to $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$. All volumes are measured at the same temperature and pressure. Determine the composition and total volume of the final mixture.

[4]

SOLUTION

At the same T and P , volume ratios equal mole ratios, so work directly in cm^3 .

50 cm^3 of N_2 would require 150 cm^3 of H_2 ; only 200 cm^3 is available, which is more than enough — so **nitrogen is limiting**. (1)

N_2 used = 50 cm^3 ; H_2 used = 150 cm^3 , leaving $200 - 150 = 50 \text{ cm}^3$ of H_2 in excess. (1)

NH_3 formed = $2 \times 50 = 100 \text{ cm}^3$. (1)

Final mixture: **$100 \text{ cm}^3 \text{ NH}_3 + 50 \text{ cm}^3 \text{ H}_2$, total 150 cm^3** . (1)

EXAMINER'S NOTE

The total volume falls from 250 cm^3 to 150 cm^3 because 4 volumes of gas become 2. Gas-syringe questions frequently ask you to explain such a decrease — the answer is the change in the total number of gas molecules, not a change in the size of the molecules.

Part G — Practice: Paper 1A multiple choice

20 QUESTIONS · SL + HL

All 20 questions are common core. Allow about 30 minutes. A calculator and the data booklet are permitted in Paper 1A under the current format.

1. What is the molar volume of an ideal gas at STP?

- A $22.4 \text{ dm}^3 \text{ mol}^{-1}$ B $22.7 \text{ dm}^3 \text{ mol}^{-1}$ C $24.0 \text{ dm}^3 \text{ mol}^{-1}$ D $24.8 \text{ dm}^3 \text{ mol}^{-1}$

2. Which conditions define STP?

- A 273 K and 101.3 kPa B 273 K and 100 kPa
C 298 K and 100 kPa D 298 K and 101.3 kPa

3. What volume is occupied by 0.500 mol of an ideal gas at STP?

- A 5.68 dm^3 B 11.4 dm^3 C 22.7 dm^3 D 45.4 dm^3

4. 1.0 dm^3 of an ideal gas at 100 kPa and 25°C is heated to 50°C at constant pressure. What is the new volume, in dm^3 ?

- A 0.50 B 0.90 C 1.1 D 2.0

5. What is the volume of gas when the pressure on 80 cm^3 of gas is changed from 300 kPa to 150 kPa at constant temperature?

- A 40.0 cm^3 B 80 cm^3 C 120 cm^3 D 160 cm^3

6. The volume of a sample of gas measured at 27°C is 5.0 dm^3 . What is the temperature when the volume is reduced, at the same pressure, to 4.0 dm^3 ?

- A -33.0°C B -3.0°C C 21.6°C D 24.3°C

7. Which graph shows the relationship between the volume and the pressure of a fixed amount of an ideal gas at constant temperature?

- | | |
|--|--|
| A a straight line with a negative gradient | B a curve that decreases, approaching but never meeting the axes |
| C a horizontal straight line | D a straight line through the origin |

8. The ideal gas equation is $PV = nRT$. Identify the units of R when pressure is measured in Pa and volume in m^3 .

- A J K^{-1} B $\text{Pa m}^3 \text{ K}^{-1} \text{ mol}^{-1}$ C J mol^{-1} D J mol K^{-1}

9. 0.0673 g of a gas was found to occupy 41.0 cm^3 at 25°C and $1.0 \times 10^5 \text{ Pa}$. What is the relative molecular mass of the gas?

- A 28 B 32 C 40 D 44

10. Which gas has the greatest density at STP?

- A H_2 B He C N_2 D CO_2

11. What amount of any ideal gas is contained in 1.00 m^3 at STP?

- A 4.41 mol B 22.7 mol C 44.1 mol D 100 mol

12. The pressure and the absolute temperature of a fixed mass of gas are both doubled. What happens to its volume?

- A It is quartered. B It is halved. C It is unchanged. D It doubles.

13. To double the volume of a fixed mass of an ideal gas at constant pressure, it is necessary to:

- A double the temperature in $^{\circ}\text{C}$ B double the temperature in K
C halve the temperature in K D double the pressure

14. Equal masses of hydrogen and oxygen are held at the same temperature and pressure. What is the ratio of the volume of hydrogen to the volume of oxygen?

- A 1 : 16 B 1 : 1 C 8 : 1 D 16 : 1

15. Equal volumes of H_2 , He and CO_2 at the same temperature and pressure contain equal:

- A masses B densities C numbers of molecules D numbers of atoms

16. What volume, at STP, is occupied by 8.0 g of oxygen, O_2 ?

- A 2.8 dm^3 B 5.7 dm^3 C 11.4 dm^3 D 22.7 dm^3

17. Which set of units may be substituted directly into $PV = nRT$ when $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$?

- A kPa, cm^3 , $^{\circ}\text{C}$ B Pa, dm^3 , K C kPa, dm^3 , K D atm, m^3 , K

18. 2.0 mol of an ideal gas is contained in 5.0 dm^3 at 300 K. What is the pressure?

- A $1.0 \times 10^3 \text{ Pa}$ B $1.0 \times 10^6 \text{ Pa}$ C $1.0 \times 10^5 \text{ Pa}$ D $2.5 \times 10^6 \text{ Pa}$

19. For a fixed mass of an ideal gas at constant pressure, the graph of volume against temperature in $^{\circ}\text{C}$ is:

- A a straight line through the origin B a straight line meeting the temperature axis at -273°C
C a curve approaching the axes D a horizontal straight line

20. **HL Stretch** What is the molar volume of an ideal gas at 100 kPa and 373 K?

- A $22.7 \text{ dm}^3 \text{ mol}^{-1}$ B $24.8 \text{ dm}^3 \text{ mol}^{-1}$ C $31.0 \text{ dm}^3 \text{ mol}^{-1}$ D $37.3 \text{ dm}^3 \text{ mol}^{-1}$

Part H — Practice: Paper 1B data-based

2 QUESTIONS · SL + HL

Question H1 — Pressure and volume of a fixed mass of gas SL + HL [12 marks]

A student used a gas syringe connected to a pressure sensor to investigate how the pressure of a fixed mass of air depends on its volume. The temperature was held at 293 K throughout.

Volume, V / cm^3	60.0	50.0	40.0	30.0	25.0	20.0
Pressure, P / kPa	100	119	149	199	238	297

- (a) Calculate the product PV for the first and last results, and state what the values show about the relationship between P and V . [3]
- (b) Explain why a graph of P against $1/V$ is a better test of the relationship than a graph of P against V . [2]
- (c) The gradient of the graph of P against $1/V$ is $5.96 \times 10^3 \text{ kPa cm}^3$. Use this value to determine the amount, in mol, of air in the syringe. [3]
- (d) State one assumption you made in part (c). [1]
- (e) The values of PV are not all identical. Suggest **two** reasons for this. [2]
- (f) Predict the volume of the gas when the pressure is 250 kPa. [1]

Question H2 — Determining a molar mass

SL + HL

HL Stretch (d)-(f)

[11

marks]

A student determined the molar mass of an unknown gas. A glass bulb of volume 155 cm^3 was evacuated, weighed, filled with the gas and reweighed. The mass of gas was found to be 0.277 g . The laboratory temperature was $22.0 \text{ }^\circ\text{C}$ and the pressure 99.5 kPa .

- (a) Calculate the amount, in mol, of gas in the bulb. [2]
- (b) Determine the molar mass of the gas. [2]
- (c) Suggest the identity of the gas. [1]
- (d) The uncertainties in the measurements were: mass $\pm 0.001 \text{ g}$, volume $\pm 1 \text{ cm}^3$, temperature $\pm 0.5 \text{ K}$, pressure $\pm 0.1 \text{ kPa}$. Calculate the total percentage uncertainty in the molar mass and identify the measurement that contributes most to it. [3]
- (e) Explain why it was necessary to evacuate the bulb before weighing it. [2]
- (f) State and explain the effect on the calculated molar mass if the gas had significant intermolecular attractions. [1]



Part I — Practice: Paper 2 structured

3 QUESTIONS · SL + HL

Question I1 — Molar volume and stoichiometry

SL + HL

[10 marks]

Calcium carbonate decomposes on strong heating: $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

- (a) State the conditions that define STP and the molar volume of an ideal gas at STP. [2]
- (b) Calculate the volume of carbon dioxide, measured at STP, produced by the complete decomposition of 2.50 g of calcium carbonate. [3]
- (c) Calculate the volume the same amount of carbon dioxide would occupy at 25.0 °C and 98.0 kPa. [3]
- (d) Explain, in terms of the ideal gas model, why the volume in (c) is larger than the volume in (b). [2]

Question I2 — A cylinder of helium

HL Stretch

[11 marks]

A helium gas cylinder has a volume of $3.00 \times 10^{-2} \text{ m}^3$. The pressure inside is $2.00 \times 10^6 \text{ Pa}$ and the temperature is 27.0 °C.

- (a) Determine the amount, in mol, of helium in the cylinder. [2]
- (b) Calculate the number of helium atoms in the cylinder. [1]
- (c) Estimate the average volume available to one helium atom. [2]
- (d) Hence estimate the average distance between neighbouring helium atoms. [2]
- (e) The radius of a helium atom is approximately $3.1 \times 10^{-11} \text{ m}$. Use your answer to (d) to comment on the validity of the assumption that the particles have negligible volume. [2]
- (f) Explain why helium behaves as an ideal gas over a wide range of pressures and temperatures. [2]

Question I3 — A weather balloon SL + HL [12 marks]

A weather balloon filled with helium has a volume of 2.00 m^3 at ground level, where the pressure is $1.01 \times 10^5 \text{ Pa}$ and the temperature is 20.0°C . It rises to an altitude where the pressure is $2.50 \times 10^4 \text{ Pa}$ and the temperature is -50.0°C .

- (a) Calculate the volume of the balloon at this altitude. [3]
- (b) Explain, in terms of the ideal gas model, why the volume changes in the way that it does, referring to **both** the change in pressure and the change in temperature. [3]
- (c) Determine the amount, in mol, of helium in the balloon. [2]
- (d) Calculate the mass of helium in the balloon. [1]
- (e) Weather balloons are deliberately filled only partly at ground level. Explain why. [2]
- (f) State, with a reason, whether the assumption of ideal behaviour is reasonable for helium under the conditions at altitude. [1]



Part J — Mark schemes

FULL SOLUTIONS

J1 Paper 1A multiple choice — answers and distractor analysis

Q	Ans	Working, and why the other options are wrong
1	B	Data booklet value at 273 K and 100 kPa. A is the obsolete value for 101.3 kPa; D is the molar volume at 298 K, not STP.
2	B	STP = 273 K and 100 kPa. 298 K belongs to "standard state" in thermodynamics; 101.3 kPa is 1 atm, no longer used as standard pressure.
3	B	$0.500 \times 22.7 = 11.35 \approx 11.4 \text{ dm}^3$. A comes from using 22.7/4; D from multiplying by 2 instead of 0.5.
4	C	Constant P : $V_2 = 1.0 \times 323/298 = 1.08 \approx 1.1 \text{ dm}^3$. D results from using 50/25 in $^{\circ}\text{C}$ — the classic kelvin error.
5	D	Constant T : $V_2 = 80 \times 300/150 = 160 \text{ cm}^3$. Halving the pressure doubles the volume; A inverts the ratio.
6	A	$T_2 = 300 \times 4.0/5.0 = 240 \text{ K} = -33.0^{\circ}\text{C}$. D comes from working in $^{\circ}\text{C}$ throughout ($27 \times 4/5 = 21.6$).
7	B	$V \propto 1/P$ gives a hyperbola asymptotic to both axes — neither variable can reach zero. A is linear, which inverse proportionality is not.
8	B	$R = PV/nT$, so the units are $\text{Pa m}^3 \text{ K}^{-1} \text{ mol}^{-1}$ (equivalent to $\text{J K}^{-1} \text{ mol}^{-1}$). D has kelvin and mole the wrong way up.
9	C	$n = PV/RT = (1.0 \times 10^5 \times 41.0 \times 10^{-6})/(8.31 \times 298) = 1.66 \times 10^{-3} \text{ mol}$; $M = 0.0673/1.66 \times 10^{-3} = 40.6 \approx 40$ (argon).
10	D	At the same T and P , $\rho = M/V_m$, so density is proportional to molar mass. CO_2 has the largest M (44.01).
11	C	$n = 1.00 / 2.27 \times 10^{-2} = 44.1 \text{ mol}$. B confuses the molar volume itself with an amount.
12	C	$P_1 V_1 / T_1 = 2 P_1 V_2 / 2 T_1$; the 2s cancel, so $V_2 = V_1$. The two changes exactly offset each other.
13	B	$V \propto T(\text{K})$ at constant pressure. A is the kelvin error: doubling from 20°C to 40°C raises T only from 293 K to 313 K.
14	D	Equal masses, say 32 g: $n(\text{H}_2) = 16 \text{ mol}$, $n(\text{O}_2) = 1 \text{ mol}$. Equal amounts occupy equal volumes, so the ratio is 16 : 1.
15	C	Avogadro's law. D fails because the gases have different numbers of atoms per molecule (1, 2 and 3).

16	B	$n = 8.0/32.0 = 0.25 \text{ mol}$; $V = 0.25 \times 22.7 = 5.68 \approx 5.7 \text{ dm}^3$. C results from using $M = 16.0$ (atomic instead of molecular oxygen).
17	C	kPa with dm^3 : the factors 10^3 and 10^{-3} cancel. B mixes Pa with dm^3 , which does not cancel; A uses $^\circ\text{C}$; D uses atm.
18	B	$P = nRT/V = (2.0 \times 8.31 \times 300)/(5.0 \times 10^{-3}) = 9.97 \times 10^5 \approx 1.0 \times 10^6 \text{ Pa}$. C arises from leaving the volume in dm^3 .
19	B	Linear but not proportional: extrapolating to $V = 0$ gives $-273 \text{ }^\circ\text{C}$. A is the graph against kelvin.
20	C	$V_m = RT/P = (8.31 \times 373)/(1.00 \times 10^5) = 3.10 \times 10^{-2} \text{ m}^3 \text{ mol}^{-1} = 31.0 \text{ dm}^3 \text{ mol}^{-1}$. Equivalently $22.7 \times 373/273$.

PATTERN TO NOTICE

Almost every wrong option in this set comes from one of four slips: (i) using $^\circ\text{C}$ instead of K; (ii) inverting a ratio; (iii) a factor-of- 10^3 or 10^6 volume conversion; (iv) applying $22.7 \text{ dm}^3 \text{ mol}^{-1}$ away from STP. Eliminate those four and the multiple-choice section on this topic becomes routine.



iBLaurel

J2 Paper 1B — Question H1 mark scheme [12]

(a) [3]

- $100 \times 60.0 = 6.00 \times 10^3 \text{ kPa cm}^3$ (1)
- $297 \times 20.0 = 5.94 \times 10^3 \text{ kPa cm}^3$ (1)
- The product is (very nearly) constant, so the pressure is inversely proportional to the volume: $P \propto 1/V$, as predicted for an ideal gas at constant temperature. (1)

(b) [2]

- P against V is a curve, and the eye cannot reliably distinguish a true hyperbola from any other decreasing curve. (1)
- P against $1/V$ should be a **straight line through the origin**; deviation from a straight line, or a non-zero intercept, is easy to detect, and the gradient has a physical meaning (nRT). (1)

(c) [3]

- gradient = nRT , so $n = \text{gradient} / RT$ (1)
- Convert: $5.96 \times 10^3 \text{ kPa cm}^3 = 5.96 \times 10^6 \text{ Pa} \times 10^{-6} \text{ m}^3 = 5.96 \text{ J}$ (1)
- $n = 5.96 / (8.31 \times 293) = 5.96 / 2435 = \mathbf{2.45 \times 10^{-3} \text{ mol}}$ (1)

A common slip is to omit the unit conversion and obtain 2.45 mol — an amount of gas that would need a 60 dm³ syringe. Award the first two marks for correct method with a conversion error.

(d) [1]

- That the air behaves as an ideal gas over this range of pressures (accept: that the temperature stayed constant at 293 K; that no gas leaked from the syringe). (1)

(e) [2]

Any **two** of:

- random uncertainty in reading the syringe volume and the pressure sensor; (1)
- the temperature may not have stayed exactly constant — compressing a gas warms it slightly, and time is needed to re-equilibrate; (1)
- a small leak, or gas trapped in the connecting tube whose volume is not counted; (1)
- air is a real gas and deviates slightly from ideal behaviour, particularly as the pressure rises. (1)

(f) [1]

- $V = PV/P = 5.96 \times 10^3 / 250 = \mathbf{23.8 \text{ cm}^3}$ (accept 23–24 cm³). (1)

J3 Paper 1B — Question H2 mark scheme [11]

(a) [2]

- $P = 99.5 \times 10^3 \text{ Pa}$, $V = 155 \times 10^{-6} \text{ m}^3$, $T = 295 \text{ K}$ (1)
- $n = PV/RT = (99.5 \times 10^3 \times 155 \times 10^{-6}) / (8.31 \times 295) = \mathbf{6.29 \times 10^{-3} \text{ mol}}$ (1)

(b) [2]

- $M = m/n = 0.277 / 6.29 \times 10^{-3}$ (1)
- = **44.0 g mol⁻¹** (1)

ECF from (a). Accept 43.9–44.1.

(c) [1]

- **Carbon dioxide**, CO₂ (accept propane, C₃H₈, or dinitrogen oxide, N₂O — all have $M \approx 44$). (1)

(d) [3]

- mass: $0.001/0.277 \times 100 = 0.36\%$; volume: $1/155 \times 100 = 0.65\%$; temperature: $0.5/295 \times 100 = 0.17\%$; pressure: $0.1/99.5 \times 100 = 0.10\%$ (1)
- total = $0.36 + 0.65 + 0.17 + 0.10 = \mathbf{1.3\%}$ (1)
- The **volume** of the bulb contributes most. (1)

Percentage uncertainties are added for quantities that are multiplied or divided. Accept 1.28–1.3 %. An absolute uncertainty in M of $\pm 0.6 \text{ g mol}^{-1}$ may also be quoted.

(e) [2]

- If the bulb still contained air, the "mass of gas" recorded would be the mass of the unknown gas *minus* the mass of the air it displaced, not the true mass. (1)
- Evacuating the bulb ensures that the increase in mass on filling is due entirely to the unknown gas, so that m and n refer to the same sample. (1)

(f) [1]

- The calculated molar mass would be slightly **too high**: attractions reduce the measured pressure below the ideal value, so n calculated from PV/RT is too small and m/n is too large. (1)

J4 Paper 2 — Question I1 mark scheme [10]

(a) [2]

- STP is 273 K (0 °C) and 100 kPa. (1)
- Molar volume at STP = $22.7 \text{ dm}^3 \text{ mol}^{-1}$ ($2.27 \times 10^{-2} \text{ m}^3 \text{ mol}^{-1}$). (1)

(b) [3]

- $M(\text{CaCO}_3) = 100.09 \text{ g mol}^{-1}$; $n = 2.50/100.09 = 0.0250 \text{ mol}$ (1)
- Mole ratio $\text{CaCO}_3 : \text{CO}_2 = 1 : 1$, so $n(\text{CO}_2) = 0.0250 \text{ mol}$ (1)
- $V = 0.0250 \times 22.7 = \mathbf{0.567 \text{ dm}^3}$ (567 cm^3) (1)

(c) [3]

- $T = 298 \text{ K}$, $P = 98.0 \times 10^3 \text{ Pa}$ (1)
- $V = nRT/P = (0.0250 \times 8.31 \times 298)/(98.0 \times 10^3)$ (1)
- $= 6.31 \times 10^{-4} \text{ m}^3 = \mathbf{0.631 \text{ dm}^3}$ (1)

Accept the alternative route via the combined gas law: $0.567 \times (298/273) \times (100/98.0) = 0.631 \text{ dm}^3$. ECF from (b).

(d) [2]

- The temperature is higher, so the particles have greater average kinetic energy and collide with the walls more frequently and more energetically; (1)
- the pressure is also lower, so the gas expands further before its collision frequency falls enough to balance the external pressure. Both changes act to increase the volume. (1)

J5 Paper 2 — Question I2 mark scheme [11]

(a) [2]

- $n = PV/RT = (2.00 \times 10^6 \times 3.00 \times 10^{-2})/(8.31 \times 300)$ (1)
- $= 60\,000 / 2493 = \mathbf{24.1 \text{ mol}}$ (1)

(b) [1]

- $24.1 \times 6.02 \times 10^{23} = \mathbf{1.45 \times 10^{25} \text{ atoms}}$ (1)

(c) [2]

- volume per atom = total volume / number of atoms = $3.00 \times 10^{-2} / 1.45 \times 10^{25}$ (1)
- $= \mathbf{2.07 \times 10^{-27} \text{ m}^3}$ (1)

(d) [2]

- Treating that volume as a cube, the separation is its cube root: $d = (2.07 \times 10^{-27})^{1/3}$ (1)
- $= \mathbf{1.3 \times 10^{-9} \text{ m}}$ (1.3 nm) (1)

(e) [2]

- The atomic diameter is about 6.2×10^{-11} m, roughly 20 times smaller than the separation, so the atoms occupy only about $(1/20)^3 \approx 0.01$ % of the available volume. (1)
- The assumption of negligible particle volume is therefore a good one even at this fairly high pressure, though it would begin to fail if the pressure were raised much further. (1)

Accept any reasoning that compares the two lengths and reaches a supported conclusion; ECF from (d).

(f) [2]

- Helium atoms are very small and monatomic, so their own volume is negligible over a wide range of conditions; (1)
- helium is non-polar with only extremely weak London dispersion forces (boiling point 4 K), so the assumption of no intermolecular forces holds well even at low temperatures. (1)



J6 Paper 2 — Question I3 mark scheme [12]

(a) [3]

- $T_1 = 293 \text{ K}$, $T_2 = 223 \text{ K}$ (1)
- $V_2 = P_1 V_1 T_2 / (T_1 P_2) = (1.01 \times 10^5 \times 2.00 \times 223) / (293 \times 2.50 \times 10^4)$ (1)
- = **6.15 m³** (1)

(b) [3]

- The external pressure falls to about a quarter of its ground-level value, so the gas expands until the reduced frequency of collisions with the inside of the balloon balances the lower external pressure. (1)
- The temperature falls, so the particles have lower average kinetic energy and collide less frequently and less energetically — this acts to *reduce* the volume. (1)
- The pressure change is much the larger effect, so the balloon expands overall, to about three times its ground-level volume. (1)

Full marks require both effects and a statement of which dominates. Two marks maximum for an answer that discusses pressure only.

(c) [2]

- $n = PV/RT = (1.01 \times 10^5 \times 2.00) / (8.31 \times 293)$ (1)
- = **83.0 mol** (1)

Either set of conditions may be used, since the amount of gas is unchanged; the altitude values give the same answer.

(d) [1]

- mass = $83.0 \times 4.00 = \mathbf{332 \text{ g}}$ (0.332 kg) (1)

(e) [2]

- The balloon expands greatly as it rises, because the external pressure falls far more than the temperature does. (1)
- Filling it completely at ground level would leave no room for that expansion and the envelope would burst at low altitude, so partial filling allows the balloon to reach the intended height. (1)

(f) [1]

- Yes — at altitude the pressure is very low, so the helium atoms are far apart and both their own volume and the (already extremely weak) intermolecular forces are negligible. The low temperature works against ideality, but for helium, whose boiling point is 4 K, 223 K is still far above the point where attractions matter. (1)

Where this topic goes next

STUDY PLAN

Related material	What it adds
Structure 1.5.1 & 1.5.2 Ideal and real gases	The assumptions behind every equation in this booklet, and the conditions under which they fail. Available as a separate iBLaurel booklet.
Structure 1.4 The mole concept	Moles, molar mass, Avogadro's law, concentration and limiting reactants — the stoichiometry that gas calculations sit on top of.
Structure 2.2 Intermolecular forces	Explains why some gases deviate more than others, and why helium and hydrogen are the closest to ideal.
Tools 2 & 3, Inquiry 2	Graphing, gradients, extrapolation, uncertainty propagation, and the experimental determination of molar mass — all assessed through gas questions.
Reactivity 1.1	Thermochemistry, where "standard state" (298 K) must be distinguished from STP (273 K).

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