

FREE REVISION RESOURCE



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Structure 1.5.1 & 1.5.2 Ideal Gases and Real Gases

The kinetic model, its assumptions and its limits

IB Diploma Programme Chemistry · First assessment 2025
Notes · Formula sheet · Exam technique · 14 solved examples · 20 MCQs · Paper 1B
& Paper 2 practice · Full mark schemes

SL + HL core content

HL Stretch harder application

Forward link 1.5.3 / 1.5.4 / 2.2

Part A — Structure 1.5.1: the ideal gas model	3
Part B — Structure 1.5.2: real gases and deviations	8
Part C — Formula, data and definitions sheet	14
Part D — Exam technique and command terms	16
Part E — Solved examples (14)	18
Part F — Practice: Paper 1A multiple choice (20)	25
Part G — Practice: Paper 1B data-based (2)	28
Part H — Practice: Paper 2 structured (3)	30
Part I — Mark schemes	31

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How to use this booklet

READ FIRST

This booklet covers the first two statements of **Structure 1.5 — Ideal gases** of the IB Diploma Programme Chemistry guide (first assessment 2025):

SYLLABUS STATEMENTS COVERED

Structure 1.5.1 — An ideal gas consists of moving particles with negligible volume and no intermolecular forces. All collisions between particles are considered elastic.

Assessment skill: Recognize the key assumptions in the ideal gas model.

Structure 1.5.2 — Real gases deviate from the ideal gas model, particularly at low temperature and high pressure.

Assessment skill: Explain the limitations of the ideal gas model.

Guidance: No mathematical coverage is required. Linking question — Structure 2.2: under comparable conditions, why do some gases deviate more from ideal behaviour than others?

What the level labels mean

Label	Meaning
SL + HL	Common core. Every SL and every HL student is examined on this. Structure 1.5 is entirely common core — there is no AHL-only (HL-only) content anywhere in Structure 1.5 . If a source tells you the ideal gas model is "HL only", it is wrong.
HL Stretch	Same syllabus content, harder question style. HL papers ask longer, more analytical versions of the same ideas — multi-step explanations, graph interpretation, comparisons between named gases. SL students who want a 7 should still work through these; nothing here is off-syllabus for SL, it is simply more demanding.
Forward link	Material that formally belongs to Structure 1.5.3 / 1.5.4 (molar volume, $PV = nRT$) or Structure 2.2 (intermolecular forces), included where a 1.5.1/1.5.2 question cannot realistically be answered without it.
BEYOND THE GUIDE	Genuinely extra (compression factor Z , van der Waals' equation). Excellent for understanding and for university preparation, but you will not be penalised for not knowing it. It is boxed in purple so you can skip it under time pressure.

THE ONE-LINE SUMMARY OF THIS WHOLE TOPIC

The ideal gas model makes gas behaviour *calculable* by throwing away two things that really exist: **particle volume** and **intermolecular forces**. Everything in Structure 1.5.2 is the story of what happens when the two things you threw away become too big to ignore — which is exactly at **high pressure** and **low temperature**.

Part A — Structure 1.5.1: The ideal gas model

SL + HL

A1 What an ideal gas is SL + HL

DEFINITION

An **ideal gas** is a hypothetical gas whose particles have **negligible volume** and **no intermolecular forces**, and whose collisions are **perfectly elastic**. It obeys $PV = nRT$ exactly, under all conditions of temperature and pressure.

"Ideal" does not mean "very pure" or "very good". It means *simplified so that the mathematics works*. No real gas is ideal; the model is a deliberate approximation, and Structure 1.5.2 is about knowing when the approximation is safe.

A2 The four assumptions (axioms) — learn these exactly SL + HL

#	Assumption	What it really means, and why it is made
1	Negligible particle volume — the particles occupy a negligible volume compared with the volume of the container.	The particles are treated as <i>point masses</i> . Therefore the volume available for movement is the full volume of the container, V . This is what lets us put the container volume straight into $PV = nRT$.
2	No intermolecular forces between particles (except at the instant of collision).	Particles neither attract nor repel one another. A particle therefore travels in a <i>straight line at constant speed</i> between collisions, and arrives at the wall with undiminished speed.
3	Continuous random motion — the particles have a <i>range</i> of speeds and move randomly; the average kinetic energy is proportional to the absolute temperature in kelvin.	"Random" means no preferred direction, so the pressure is the same on every wall. Note carefully: particles do <i>not</i> all move at the same speed — they have a distribution of speeds (Figure 3).
4	Elastic collisions — collisions with the walls and with each other conserve kinetic energy.	No kinetic energy is lost to the walls or converted to other forms. Without this the gas would gradually "run down" and the pressure would fall to zero on its own.

MEMORY AID — V · F · R · E

Volume negligible · Forces absent · Random motion with $E_k \propto T(K)$ · Elastic collisions. Two of these (V and F) are the ones that *fail* for real gases — the other two (R and E) hold up well. That single fact is the bridge from 1.5.1 to 1.5.2.

COMMON ERROR — "NO FORCES AT ALL"

Assumption 2 says there are **no intermolecular forces**. It does **not** say there are no forces of any kind — collisions themselves involve contact forces, and that is how the gas exerts pressure. Write "no intermolecular forces *between* the particles", not "no forces act".

A3 Where gas pressure comes from SL + HL

DEFINITION

Pressure is force per unit area. The pressure of a gas arises from the **collisions of particles with the walls of the container**. Each collision delivers a tiny impulse; a very large number of collisions per second produces a steady, measurable pressure.

Two factors, and only two, control the pressure:

1. the **frequency** of collisions with the wall (how many per second), and
2. the **force of each collision** (how energetic each impact is).

Because the motion is random (assumption 3), there is no preferred direction and so the pressure is **the same on every wall**. Because the collisions are elastic (assumption 4), a particle rebounds with the same speed it arrived with, so the pressure does not decay with time.

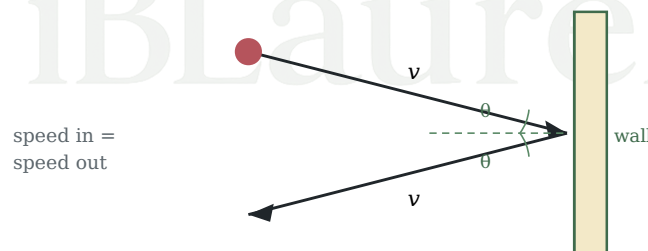


Figure 1 An elastic collision with the wall. The particle rebounds with the *same speed*, so kinetic energy is conserved. The impulse delivered to the wall is the origin of gas pressure.

A4 Using the model to explain the gas laws SL + HL

This is the single most examined skill in 1.5.1. You must be able to *explain*, in terms of collisions, why pressure changes. Always argue through **collision frequency** and/or **collision energy**.

Change	Effect on P	Explanation in terms of the model (the marks are here)
Volume increased (T, n constant)	Decreases	Particles must travel a greater distance between collisions with the walls, so they collide with the walls less frequently . Speed and energy of each collision are <i>unchanged</i> because T is unchanged. Fewer impacts per second = lower pressure. ($P \propto 1/V$)

Temperature increased (V, n constant)	Increases	Two reasons — quote both: (i) higher average kinetic energy = particles move faster $\Rightarrow$ each collision is more energetic ; (ii) moving faster, they reach the walls more often $\Rightarrow$ collisions are more frequent . ($P \propto T$)
Temperature increased (P, n constant)	— (V rises)	Each particle now collides with more energy, so the only way to keep the pressure constant is to reduce the collision frequency — which happens if the volume increases. ($V \propto T$)
More gas added (T, V constant)	Increases	More particles in the same volume $\Rightarrow$ more collisions with the walls per second. Energy per collision unchanged. ($P \propto n$)
Different gas, same n, V, T	No change	The model contains nothing about chemical identity. Heavier particles move more slowly but hit harder; the two effects cancel exactly. This is why one equation works for all gases.

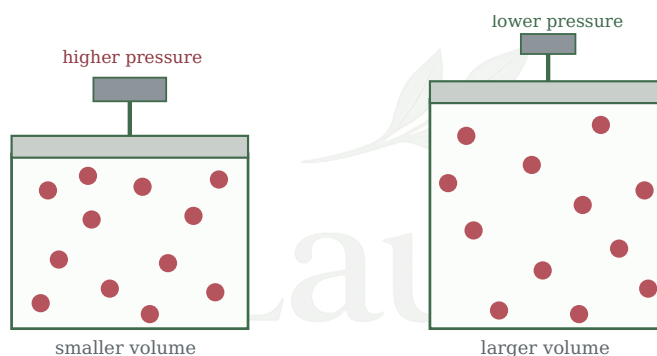


Figure 2 Same amount of gas, same temperature. In the larger volume the particles travel further between wall collisions, so collisions are *less frequent* and the pressure is lower.

A5 Temperature, kinetic energy and the speed distribution SL + HL

$$E_k(\text{average}) \propto T \text{ (in kelvin)}$$

Absolute temperature is a measure of the *average* kinetic energy of the particles (Structure 1.1.3).

- Temperature in these statements always means **absolute temperature in kelvin**. $T(\text{K}) = \theta(^{\circ}\text{C}) + 273$.
- Doubling the kelvin temperature doubles the average kinetic energy. Doubling the Celsius temperature does *not* — this is the single most common numerical error in the whole of Structure 1.5.
- At **0 K** (absolute zero) the particles would have no kinetic energy and the pressure would be zero. Absolute zero is unattainable in practice.
- Particles have a **range** of speeds, not one speed. Raising the temperature broadens and flattens the distribution and shifts it to higher speeds (Figure 3). Two gases at the same

temperature have the *same average kinetic energy*, but the lighter gas has the higher average *speed* (because $E_k = \frac{1}{2}mv^2$).

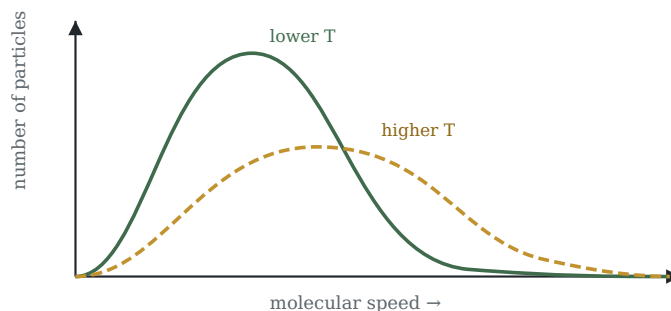


Figure 3 The distribution of molecular speeds. Assumption 3 says particles have a *range* of speeds — the average, not the individual value, is fixed by the temperature.

COMMON ERROR — "ALL THE PARTICLES MOVE AT THE SAME SPEED"

Examiners test this deliberately. A statement such as "the particles all have the same kinetic energy" is **wrong**. The correct wording is "the *average* kinetic energy of the particles is proportional to the absolute temperature".

A6 Units — get these right before you calculate anything [Forward link 1.5.4](#)

Quantity	SI unit	Conversions and traps
Pressure, P	pascal, Pa (= N m^{-2})	$1 \text{ kPa} = 10^3 \text{ Pa}$ · $1 \text{ atm} = 1.013 \times 10^5 \text{ Pa}$ · $1 \text{ bar} = 10^5 \text{ Pa}$. IB examinations use SI units; you will not be asked to convert from mmHg, torr or psi.
Volume, V	cubic metre, m^3	$1 \text{ dm}^3 = 10^{-3} \text{ m}^3$ · $1 \text{ cm}^3 = 10^{-6} \text{ m}^3$ · $1 \text{ cm}^3 = 10^{-3} \text{ dm}^3$. $\text{cm}^3 \rightarrow \text{dm}^3 \rightarrow \text{m}^3$: divide by 1000 each time. $\text{dm}^3 \equiv \text{L}$, $\text{cm}^3 \equiv \text{mL}$.
Temperature, T	kelvin, K	$T(\text{K}) = \theta(^{\circ}\text{C}) + 273$ (or 273.15 for precise work). Never substitute $^{\circ}\text{C}$. A <i>change</i> of 1°C equals a change of 1 K — the size of the degree is identical.
Amount, n	mole, mol	$n = m/M$. Watch for mass given in mg or kg.

EXAM TIP — THE "KPA AND DM^3 SHORTCUT"

If pressure is in kPa *and* volume in dm^3 , the factors of 10^3 and 10^{-3} cancel, so you may substitute those numbers directly into $PV = nRT$ and still get the right answer. This is a genuine shortcut, but only for that exact pairing. Anything else — convert to Pa and m^3 first.

A7 What the model gets right SL + HL

- It explains all three experimental gas laws ($P \propto 1/V$, $V \propto T$, $P \propto T$) from a single set of assumptions.
- It explains why the chemical identity of the gas makes no difference (Avogadro's law, Structure 1.4.6) — one equation for every gas.
- Its predictions are accurate to within a few per cent for most gases at **ordinary laboratory temperature and pressure**, which is why it is used at all.
- It is a good illustration of the nature of science: a model built on unobservable particles, tested against, and supported by, measurable experimental laws.



Part B — Structure 1.5.2: Real gases

SL + HL

DEFINITION

A **real gas** is any actual gas. Real gas particles have a **finite, measurable volume** and **do experience intermolecular forces**. Real gases therefore **deviate** from ideal behaviour, and the deviation is greatest at **high pressure** and **low temperature**.

B1 Exactly which assumptions fail SL + HL

Only assumptions **1** and **2** break down. Assumptions 3 and 4 (random motion, elastic collisions) remain good descriptions of real gases. If a question asks you to state the limitations of the model, name the two failing assumptions and say what each does to the pressure.

	Assumption that fails	Physical consequence	Effect on pressure / volume
1	Particle volume is not negligible	The particles themselves take up space, so the free space available for movement is less than the container volume. Particles travel a shorter distance between wall collisions.	Collisions are more frequent than the model predicts $\Rightarrow$ the observed pressure is greater than ideal. Equivalently, the gas is harder to compress than predicted.
2	There are attractive intermolecular forces	A particle heading for the wall is pulled backwards by its neighbours behind it, so it arrives with a reduced speed.	Collisions are less energetic $\Rightarrow$ the observed pressure is lower than ideal. Equivalently, the gas is easier to compress than predicted.

THE TWO EFFECTS OPPOSE EACH OTHER

Attractions **lower** the pressure; finite particle volume **raises** it. Which one wins depends on the conditions. At **moderate pressures** attractions dominate; at **very high pressures** the finite volume dominates. This is the whole shape of the deviation graphs in section B4.

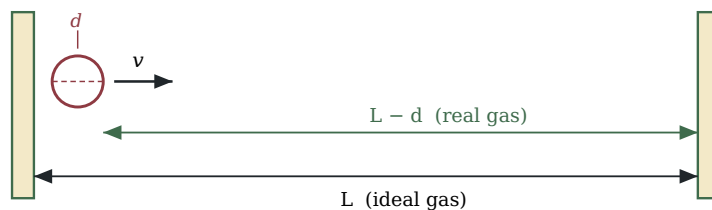


Figure 4 Assumption 1 fails. An ideal (point) particle travels a distance $2L$ between successive collisions with the same wall; a real particle of diameter d travels only $2(L - d)$. The shortfall matters when L is small — that is, at high pressure and low temperature.

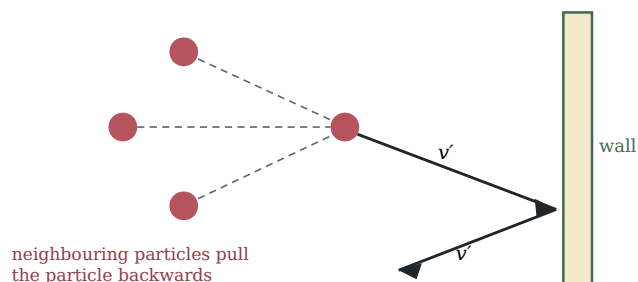


Figure 5 Assumption 2 fails. Attractions from the particles behind slow the particle as it approaches the wall, so the collision is less energetic and the measured pressure is lower than the ideal value.

B2 Why high pressure and low temperature? SL + HL

Condition	Why the model fails there
High pressure	The gas is compressed into a small volume, so the particles are close together . (i) The particles' own volume becomes a significant <i>fraction</i> of the container volume — assumption 1 fails. (ii) Being close together, intermolecular forces act strongly — assumption 2 fails.
Low temperature	The particles move more slowly (lower average E_k), so they spend longer near one another and the attractive forces have more time and more effect on their motion. The forces are no longer negligible compared with the kinetic energy — assumption 2 fails.
Ideal behaviour is approached at: low pressure and high temperature	Low pressure $\Rightarrow$ particles far apart $\Rightarrow$ their own volume is negligible and forces are weak. High temperature $\Rightarrow$ kinetic energy is large compared with the intermolecular attractions, so the attractions are effectively overwhelmed. A useful rule of thumb: a gas behaves most ideally when it is far above its boiling point and at low density .

EXAM TIP — THE INTUITION THAT NEVER FAILS YOU

Deviation from ideality is largest when the gas is **closest to condensing into a liquid**. High pressure and low temperature are precisely the conditions used to liquefy a gas. So: "*the nearer the gas is to becoming a liquid, the less ideally it behaves.*" If you can only remember one sentence about 1.5.2, remember that one.

B3 Which gases deviate most?

[Forward link — Structure 2.2](#)
[HL Stretch](#)

Under the *same* conditions, gases differ in how far they deviate. Two molecular properties decide it:

Property	Effect on deviation
Strength of intermolecular forces (hydrogen bonding > dipole-dipole > London dispersion)	The stronger the forces, the more the particles are pulled back from the wall, and the greater the deviation (pressure lower than ideal). Polar molecules and hydrogen-bonded molecules deviate most.
Size / molar mass of the particles	Larger particles occupy a larger fraction of the container and also have more electrons, giving stronger London dispersion forces. Both effects increase deviation.

Ranking examples you should be able to justify:

- **NH₃ deviates more than CH₄** — ammonia is polar and hydrogen bonds; methane is non-polar with only weak London forces. Similar molar masses (17 and 16 g mol⁻¹), so intermolecular forces are the deciding factor.
- **H₂O(g) deviates more than H₂S(g)** — hydrogen bonding beats dipole-dipole, despite water's smaller molar mass.
- **He and H₂ are the closest to ideal** — very small particles, very few electrons, extremely weak London forces, and very low boiling points (4 K and 20 K), so at room temperature they are hundreds of kelvin above their boiling points.
- **Cl₂ deviates more than He** — much larger and far more polarisable, with much stronger London forces.
- Among gases with only London forces, deviation increases down a group: He < Ne < Ar < Kr < Xe.

COMMON ERROR — "POLAR GASES DEVIATE BECAUSE THEY ARE HEAVIER"

Give the *reason*, not the correlation. The mark is for naming the **type** of intermolecular force (hydrogen bonding, dipole-dipole, London dispersion) and comparing its strength, or for referring to particle size/polarisability — not for quoting M_r alone.

B4 Reading deviation graphs

[HL Stretch](#)

Deviation is usually shown by plotting PV/nRT (for 1 mol, PV/RT) against pressure. For an ideal gas this quantity equals exactly **1** at all pressures, giving a horizontal line — the reference against which real gases are compared.

$$PV / nRT = 1 \quad \text{for an ideal gas, at all } P \text{ and } T$$

Above 1 = particle volume dominating. Below 1 = intermolecular attractions dominating.

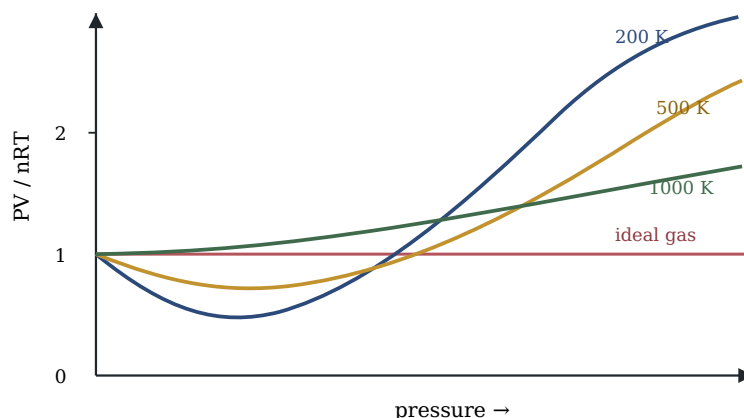


Figure 6 Deviation of a real gas (nitrogen) at three temperatures. The dip below 1 is caused by intermolecular attractions; the rise above 1 at high pressure is caused by the non-negligible volume of the particles. Deviation is smallest at the highest temperature and at the lowest pressure.

How to read Figure 6 in an exam

- **All curves start at 1** as $P \rightarrow 0$. At vanishing pressure every gas behaves ideally, because the particles are infinitely far apart.
- **Dip below 1** (moderate pressure) — attractive forces are pulling particles back from the walls, so the measured PV is smaller than nRT . The deeper the dip, the stronger the intermolecular forces.
- **Rise above 1** (high pressure) — the particles' own volume is now a large fraction of the container, so the measured volume is larger than an ideal gas would occupy and PV exceeds nRT . The steeper the rise, the larger the particles.
- **Raising the temperature flattens the curve** towards the ideal line — at 1000 K there is barely a dip at all, because the kinetic energy of the particles overwhelms the attractions.

EXAM TIP — COMPARING SEVERAL GASES ON ONE AXIS

If three gases are plotted at the *same* temperature: the one with the **deepest dip below 1** has the **strongest intermolecular forces**; the one that **rises above 1 earliest and most steeply**, without dipping, has the **largest particle volume** and negligible attractions (typically He or H_2). Say which feature of the curve you are using — that is where the "justify your answer" mark lives.

BEYOND THE GUIDE — COMPRESSION FACTOR AND VAN DER WAALS' EQUATION

The quantity $Z = PV/nRT$ is called the **compression factor**. $Z = 1$ for an ideal gas; $Z < 1$ when attractions dominate; $Z > 1$ when finite molecular volume dominates.

The **van der Waals' equation** corrects both faults of the ideal model at once:

$$(P + an^2/V^2)(V - nb) = nRT$$

The constant **a** measures the strength of the intermolecular attractions (it *adds* back the pressure lost to attractions); the constant **b** is the volume excluded by one mole of particles (it *subtracts* the space the particles themselves occupy). Both are specific to each gas. Large **a** $\Rightarrow$ strongly polar or hydrogen-bonded gas; large **b** $\Rightarrow$ physically large molecules. **You will not be examined on this equation**, but it makes the two effects concrete, and it explains why one equation cannot fit all gases once you leave ideal conditions.

B5 Is the model therefore useless? SL + HL

No — and this is a favourite "nature of science" question. The ideal gas model is:

- **accurate enough** for most gases at ordinary temperatures and pressures (errors of a few per cent at most);
- **universal** — a single equation covers every gas, whereas van der Waals' equation needs two experimentally determined constants for each gas;
- **a limiting case** — every real gas tends towards ideal behaviour as pressure tends to zero, so the model is the correct starting point rather than a wrong answer.

A model is judged by whether it is *useful within a stated range*, not by whether it is literally true. Knowing the limits of a model is part of understanding it.

B6 Summary comparison you can reproduce in 90 seconds SL + HL

Feature	Ideal gas	Real gas
Particle volume	Negligible (point masses)	Finite and measurable
Intermolecular forces	None	Present; strength depends on the substance
Collisions	Perfectly elastic	Essentially elastic — this assumption holds well
Motion	Continuous, random, straight-line	The same — this assumption holds well
Obeys $PV = nRT$	Exactly, under all conditions	Only approximately; exactly only in the limit $P \rightarrow 0$
PV/nRT	Exactly 1 always	< 1 (attractions dominate) or > 1 (volume dominates)

Can be liquefied?	Never — no attractions, so it could never condense	Yes, on cooling and compression
Behaves most ideally	—	At high temperature and low pressure ; small, non-polar particles (He, H ₂ , Ne)



Part C — Formula, data and definitions sheet

ONE-PAGE REVISION

C1 Given to you in the IB data booklet

Item	Value / form	Where and when
Ideal gas equation	$PV = nRT$	Section 1. Used for any single set of conditions.
Combined gas law	$P_1 V_1 / T_1 = P_2 V_2 / T_2$	Section 1. Used when a fixed mass of gas changes conditions.
Gas constant, R	$8.31 \text{ J K}^{-1} \text{ mol}^{-1}$	Section 2. Equivalently $8.31 \text{ N m K}^{-1} \text{ mol}^{-1}$.
Molar volume at STP	$2.27 \times 10^{-2} \text{ m}^3 \text{ mol}^{-1}$ ($= 22.7 \text{ dm}^3 \text{ mol}^{-1}$)	Section 2. STP = 273 K and 100 kPa. Not 22.4 — that is the obsolete value for 101.3 kPa.
Avogadro's constant	$6.02 \times 10^{23} \text{ mol}^{-1}$	Section 2.

C2 Relationships you must know without the booklet

Relationship	Conditions and meaning
$P \propto 1/V$ ($PV = \text{constant}$)	Fixed n , constant T . Graph of P against V is a curve; P against $1/V$ is a straight line through the origin; PV against P is horizontal.
$V \propto T(\text{K})$ ($V/T = \text{constant}$)	Fixed n , constant P . Straight line through the origin when T is in kelvin; straight line with a positive intercept when plotted against $^{\circ}\text{C}$, crossing the temperature axis at -273°C .
$P \propto T(\text{K})$ ($P/T = \text{constant}$)	Fixed n , constant V . Same graph shapes as above.
$E_k(\text{average}) \propto T(\text{K})$	Absolute temperature measures average kinetic energy. Same $T =$ same average E_k for every gas.
$T(\text{K}) = \theta(^{\circ}\text{C}) + 273$	Always convert before substituting.
$n = m/M$ and $M = mRT/PV$	The second is the ideal gas equation rearranged for molar mass — the standard route from experimental data to M . 1.5.4
density $\rho = m/V$, so $M = \rho RT/P$	Used when the data are given as a gas density. 1.5.4
$Z = PV/nRT$	Compression factor. = 1 ideal; < 1 attractions dominate; > 1 particle volume dominates. HL Stretch

C3 Definitions in examiner-friendly wording

Term	Answer that scores
Ideal gas	A gas whose particles have negligible volume and no intermolecular forces, undergo perfectly elastic collisions, and obey $PV = nRT$ under all conditions.
Real gas	An actual gas, whose particles have finite volume and experience intermolecular forces, so it deviates from $PV = nRT$, especially at high pressure and low temperature.
Elastic collision	A collision in which the total kinetic energy is conserved — the particle rebounds with the same speed.
Pressure of a gas	The force per unit area exerted on the container walls by the collisions of the gas particles.
Absolute zero	0 K (−273 °C): the temperature at which the particles would have no kinetic energy and the gas would exert no pressure; unattainable in practice.
STP	Standard temperature and pressure: 273 K and 100 kPa. (Not the same as "standard state" in Reactivity 1.1.)

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Part D — Exam technique

SL + HL

D1 Command terms in this topic, and what each demands

Command term	What the examiner wants	Worked phrasing
State	A short answer, no reasoning.	"The particles have negligible volume."
Identify	Pick the correct item from those given.	"Assumption 2 — no intermolecular forces."
Outline / Describe	A brief account, no justification needed.	"Pressure increases as temperature rises."
Explain	Reasons or causes. This is where you must argue through collision frequency and/or collision energy .	"Pressure increases because the particles have greater average kinetic energy, so they collide with the walls more frequently and more energetically."
Justify	Give evidence for a choice you have already made — usually a feature of a graph.	"Gas III, because its curve dips furthest below 1, showing the strongest intermolecular attractions."
Deduce	Reach a conclusion from the information given.	Reading the ranking of gases straight off a deviation graph.
Comment on	Judge the quality of data or of a model against evidence.	"The line is straight, so the gas behaves ideally over this range."

D2 Twelve habits that turn a 5 into a 7

#	Habit
1	Convert to kelvin first, every single time. Write $T = \dots \text{K}$ at the top of your working before anything else. This one habit recovers more marks in Structure 1.5 than any other.
2	Explain pressure only through frequency and energy of collisions. If your sentence contains neither word, it is unlikely to score.

3	When temperature changes, quote both effects (more frequent <i>and</i> more energetic collisions). When only volume changes at constant T , quote frequency only — and say explicitly that the speed is unchanged.
4	Never say the particles "expand" or "get bigger" . Particles do not change size. The gas expands; the particles do not.
5	Never say intermolecular forces "decrease" when the gas expands in an <i>ideal</i> gas question — in the ideal model they are zero to begin with. This is a standard MCQ distractor.
6	Name the failing assumption when explaining deviation, then say whether the pressure comes out higher or lower than ideal. Two ideas = two marks.
7	Give the type of intermolecular force by name (hydrogen bonding / dipole-dipole / London dispersion) when comparing gases. "Stronger forces" alone is usually only half the mark.
8	Use the graph, not memory, in graph questions. Quote the feature: "dips below 1", "rises above 1 above 4×10^7 Pa", "straight line through the origin".
9	Check what is being held constant before choosing a relationship. Most gas-law errors are actually control-variable errors.
10	Watch significant figures. Give the answer to the same number of significant figures as the least precise datum, usually 3. Keep full precision in your calculator until the final step.
11	Sanity-check the direction. Heating at fixed volume must raise the pressure; compressing must raise it. If your answer moves the wrong way, you have probably used $^{\circ}\text{C}$ or inverted a ratio.
12	Answer the level of the question. A 1-mark "state" needs one clause; a 3-mark "explain" needs three distinct scoring points, not one point written three ways.

A 30-SECOND SELF-TEST BEFORE THE EXAM

Can you, from memory: (1) list the four assumptions; (2) say which two fail and what each does to the pressure; (3) state the two conditions of greatest deviation and the two of most ideal behaviour; (4) explain why NH_3 deviates more than CH_4 ; (5) sketch PV/nRT against P for an ideal gas and for a real gas? If yes, you have Structure 1.5.1 and 1.5.2 secured.

Part E — Solved examples

14 WORKED ANSWERS

Each example shows the answer an examiner would award full marks to, plus the reasoning that produced it. Work through the question yourself before reading the solution.

Example 1 · Stating the model

SL + HL

State the four assumptions of the ideal gas model.

[4]

SOLUTION

1. The particles occupy a negligible volume compared with the volume of the container.
2. There are no intermolecular forces between the particles (except during collisions).
3. The particles are in continuous random motion with a range of speeds; the average kinetic energy is proportional to the absolute temperature.
4. All collisions (with the walls and with each other) are perfectly elastic — kinetic energy is conserved.

EXAMINER'S NOTE

Four marks means four *distinct* ideas. "The particles are small and move fast" is one idea badly expressed, not two. Use the word *negligible*, and use *average* in point 3.

Example 2 · Heating at constant volume

SL + HL

Using the ideal gas model, explain why the pressure of a fixed mass of gas increases when it is heated in a sealed rigid container.

[3]

SOLUTION

- Raising the temperature increases the average kinetic energy of the particles, so they move faster. (1)
- Faster particles reach the walls more often, so the **frequency** of collisions with the walls increases. (1)
- Each collision is also more **energetic**, so a greater force is exerted per collision. Both effects raise the force per unit area, i.e. the pressure. (1)

EXAMINER'S NOTE

Three marks, three ideas: energy up → frequency up → force per impact up. Candidates who write only "the particles move faster so the pressure increases" score one.

Example 3 · Expansion at constant temperature

SL + HL

The volume of an ideal gas in a container is increased at constant temperature. State and explain the effect on (a) the speed of the particles and (b) the pressure.

[3]

SOLUTION

(a) No change. The average kinetic energy — and therefore the average speed — depends only on the absolute temperature, which is constant. (1)

(b) The pressure decreases. (1) The particles now travel greater distances between collisions with the walls, so the frequency of wall collisions falls; with the force per collision unchanged, the force per unit area falls. (1)

EXAMINER'S NOTE

The trap here is to attribute the drop in pressure to "weaker intermolecular forces". In an *ideal* gas the intermolecular forces are zero at every volume, so they cannot decrease. This exact distractor appears in Paper 1 year after year.

Example 4 · Monatomic gases

SL + HL

Identify which of the four assumptions of the ideal gas model apply to a monatomic gas such as argon. [2]

SOLUTION

All four apply. The assumptions are stated in terms of "particles", and an atom is as valid a particle as a molecule. (1)

Indeed a monatomic noble gas is one of the *best* approximations to an ideal gas: the atoms are small, non-polar and have only weak London dispersion forces, so assumptions 1 and 2 are unusually well satisfied. (1)

EXAMINER'S NOTE

Watch for wording traps: an assumption phrased as "there are no forces between the *molecules*" still applies — read "molecule" as "particle". Do not claim that monatomic gases are exempt from some assumption because they have no bonds.

Example 5 · If collisions were inelastic

HL Stretch

Suggest what would happen to the pressure of a gas in a sealed, thermally insulated container if the collisions between its particles were *not* perfectly elastic. [3]

SOLUTION

- Kinetic energy would be lost at each collision (converted into other forms, e.g. vibrational or internal energy). (1)
- The average kinetic energy — and therefore the average speed — of the particles would steadily fall, so collisions with the walls would become both less frequent and less energetic. (1)
- The pressure would therefore fall progressively with time, eventually approaching zero: a gas would spontaneously "run down" and could not maintain a steady pressure. Since real gases plainly *do* maintain constant pressure, collisions must be essentially elastic. (1)

EXAMINER'S NOTE

"Suggest" invites reasoning from the model, not recall. The strongest answers close the loop by pointing out that the observed constancy of pressure is the evidence *for* assumption 4.

Example 6 · Deviation at low temperature
SL + HL

Explain the difference in behaviour between a real gas and an ideal gas at low temperature.

[3]

SOLUTION

- At low temperature the particles have a low average kinetic energy and therefore move slowly. (1)
- The attractive intermolecular forces — assumed to be zero in the ideal model — are no longer negligible compared with the kinetic energy, and act on the particles for longer. (1)
- Particles approaching a wall are pulled back by their neighbours, so they collide with less energy and the measured pressure is **lower** than the ideal gas equation predicts. (1)

EXAMINER'S NOTE

Notice the structure: *slow particles* → *forces matter* → *less energetic wall collisions* → *lower pressure*. Marks are awarded for the chain, not for the conclusion alone.

Example 7 · The conditions for ideal behaviour
SL + HL

Under what conditions of temperature and pressure are real gases most likely to behave like ideal gases? Justify your answer.

[3]

SOLUTION

High temperature and low pressure. (1)

- At **low pressure** the particles are far apart, so their own volume is a negligible fraction of the container volume and the intermolecular forces between them are very weak (assumptions 1 and 2 are both good). (1)
- At **high temperature** the particles have a high average kinetic energy, which is large compared with the energy of the intermolecular attractions, so those attractions have little effect on the motion. (1)

EXAMINER'S NOTE

A useful extra sentence for a "comment" or "discuss" question: these are exactly the conditions *furthest* from those used to liquefy a gas.

Example 8 · Comparing two gases
Structure 2.2
HL Stretch

Ammonia, NH_3 , forms hydrogen bonds; methane, CH_4 , does not. Explain the relative deviation from ideal behaviour that each gas is likely to show under the same conditions. [3]

SOLUTION

- NH_3 is polar and its molecules hydrogen bond to one another, whereas CH_4 is non-polar with only weak London dispersion forces. (1)
- The intermolecular attractions in ammonia are therefore much stronger, and assumption 2 of the ideal model is a much poorer approximation for NH_3 . (1)
- Ammonia therefore deviates more from ideal behaviour: its molecules are slowed more on approaching the walls, so its pressure falls further below the ideal value (PV/nRT dips further below 1). (1)

EXAMINER'S NOTE

Their molar masses are almost identical (17 and 16 g mol^{-1}), which is exactly why the examiner chose this pair — you cannot explain the difference by size, only by the type of intermolecular force. Ammonia's boiling point (-33°C) is also far higher than methane's (-162°C), which is independent evidence of the stronger forces.

Example 9 · Reading a deviation graph

HL Stretch

Three gases were investigated over a range of pressures at the same temperature. Curve I rises steadily above 1 from the start; curve II is the horizontal ideal line; curve III dips steeply to a minimum of about 0.45 before rising and crossing 1 at high pressure.

- (a) Identify the gas with the strongest intermolecular forces and justify your answer. [2]
 (b) Identify the gas with the largest molecular volume and justify your answer. [2]

SOLUTION

(a) Gas III. (1) Its curve dips furthest below $PV/nRT = 1$, which shows that its measured pressure is well below the ideal value; this depression is caused by attractive intermolecular forces pulling particles back from the walls, so gas III must have the strongest attractions. (1)

(b) Gas I. (1) It lies above 1 at all pressures and rises most steeply, with no dip at all; values above 1 arise when the particles' own volume is not negligible, so gas I has the largest molecular volume (and negligible attractions). (1)

EXAMINER'S NOTE

You must *quote the feature of the curve* — "dips below 1", "rises above 1" — to earn the justification mark. Naming the gas alone is worth one mark at most.

Example 10 · Calculating and interpreting a compression factor

1.5.4
HL Stretch

1.00 mol of carbon dioxide occupies 1.50 dm^3 at 300 K and a pressure of $1.50 \times 10^6 \text{ Pa}$.

- (a) Calculate the volume that 1.00 mol of an ideal gas would occupy under the same conditions.
 (b) Calculate PV/nRT for the carbon dioxide and comment on which assumption of the ideal gas model is failing. [4]

SOLUTION

(a) $V = nRT/P = (1.00 \times 8.31 \times 300) / (1.50 \times 10^6) = 1.662 \times 10^{-3} \text{ m}^3 = \mathbf{1.66 \text{ dm}^3} \text{ (1)}$

(b) Convert: $V = 1.50 \text{ dm}^3 = 1.50 \times 10^{-3} \text{ m}^3$.

$$PV/nRT = (1.50 \times 10^6 \times 1.50 \times 10^{-3}) / (1.00 \times 8.31 \times 300) = 2250 / 2493 = \mathbf{0.903} \text{ (1)}$$

- The value is **less than 1**: the real gas occupies a smaller volume (exerts a lower pressure) than an ideal gas would. (1)
- This is caused by the failure of assumption 2 — CO_2 molecules do attract one another, and these attractions dominate at this pressure. (1)

EXAMINER'S NOTE

Check the direction of your conclusion against physical sense: attractions pull the gas *in*, so the real volume is *smaller*, so the ratio is *below* 1. If your arithmetic gives a value above 1 here, you have probably left the volume in dm^3 .

Example 11 · Showing that particle volume is (or is not) negligible
HL Stretch

The 18.0 g of water in 1 mol of steam would occupy about 18 cm^3 as a liquid — a reasonable estimate of the volume of the molecules themselves.

(a) Calculate the volume of 1.00 mol of steam at 500 K and $1.00 \times 10^5 \text{ Pa}$, and show that the volume of the particles is negligible.

(b) Repeat for 400 K and $5.00 \times 10^7 \text{ Pa}$, and comment. [5]

SOLUTION

(a) $V = nRT/P = (1.00 \times 8.31 \times 500) / (1.00 \times 10^5) = 4.155 \times 10^{-2} \text{ m}^3 = 41.55 \text{ dm}^3 = 41\,550 \text{ cm}^3 \text{ (1)}$

Fraction occupied by the particles = $18 / 41\,550 = 4.3 \times 10^{-4} \approx \mathbf{0.04 \%}$ — negligible, so assumption 1 is an excellent approximation here. (1)

(b) $V = (1.00 \times 8.31 \times 400) / (5.00 \times 10^7) = 6.648 \times 10^{-5} \text{ m}^3 = \mathbf{66.5 \text{ cm}^3} \text{ (1)}$

Fraction occupied = $18 / 66.5 = 0.271$, i.e. about **27 %** of the total volume. (1)

This is very far from negligible, so assumption 1 fails badly and the ideal gas equation cannot be trusted at this pressure and temperature. (1)

EXAMINER'S NOTE

This calculation is the quantitative backbone of the whole of 1.5.2: the *same* molecules are negligible at one set of conditions and dominant at another. Nothing about the gas changed — only the space it was given.

Example 12 · Why identity does not matter
1.4.6
SL + HL

Explain, using the ideal gas model, why equal volumes of any two gases at the same temperature and pressure contain equal numbers of particles. [3]

SOLUTION

- In the ideal model the particles have negligible volume, so the volume occupied depends on the *number* of particles and the space they move in, not on how big or heavy the particles are. (1)
- There are no intermolecular forces, so there is nothing that depends on the chemical identity of the gas. (1)
- At the same temperature every gas has the same average kinetic energy; heavier particles move more slowly but strike the wall harder, and these effects cancel — so the same number of particles gives the same pressure. This is Avogadro's law. (1)

Example 13 · A conceptual test of the model

HL Stretch

Explain why a truly ideal gas could never be liquefied, however much it were cooled or compressed. [2]

SOLUTION

- Condensation requires attractive intermolecular forces to hold the particles together in the liquid state. (1)
- The ideal gas model assumes there are **no** intermolecular forces, so there is nothing to hold the particles together no matter how slowly they move or how close they are pushed — the gas would remain a gas at all temperatures and pressures. (1)

EXAMINER'S NOTE

The fact that every real gas *can* be liquefied is itself proof that assumption 2 is false for all real gases. This is a strong closing sentence for any "evaluate the model" question.

Example 14 · Testing ideality from experimental data

Tools 2/3

HL Stretch

A student measured the pressure of a fixed mass of gas at constant volume over the range 5–80 °C. The line of best fit was $P = 0.3338\theta + 95.625$, with P in kPa and θ in °C.

- Describe the relationship between P and θ . [1]
- Use the equation to predict the temperature at which the pressure would be zero. [1]
- Comment on the accuracy of the student's results. [2]
- State the relationship expected between P and T in kelvin. [1]

SOLUTION

(a) A **linear** relationship with a positive gradient and a positive intercept — pressure increases uniformly with Celsius temperature, but is *not* proportional to it. (1)

(b) Set $P = 0$: $0 = 0.3338\theta + 95.625$, so $\theta = -95.625 / 0.3338 = -286$ °C (1)

(c) The extrapolated value should be absolute zero, -273 °C. The result is about 13 °C too low, an error of roughly 5 %. (1) This suggests a small **systematic** error — for example an uncalibrated pressure gauge, a small leak, or gas in the connecting tube that is not at the

measured temperature. The high linearity of the data shows the random error is small, so the fault is in accuracy rather than precision. (1)

(d) Plotted against absolute temperature, the graph would be a straight line **through the origin**: $P \propto T(\text{K})$, i.e. $P/T = \text{constant}$. (1)

EXAMINER'S NOTE

Extrapolating a P - θ or V - θ line back to the temperature axis is the classic experimental determination of absolute zero. Expect to be asked to comment on how close the value comes to $-273\text{ }^{\circ}\text{C}$, and to distinguish *systematic* from *random* error when you do.



Part F — Practice: Paper 1A multiple choice

20 QUESTIONS · SL + HL

All 20 questions are common core, so both SL and HL candidates should attempt every one. Questions marked **HL Stretch** are written in the harder style typical of HL Paper 1A. Allow about 25 minutes. A calculator and the data booklet are permitted in Paper 1A under the current format, but almost nothing here needs one.

1. Identify an assumption of the ideal gas model.

- | | |
|---|---|
| A Molecules have zero mass. | B The forces between the gas molecules and the container are always zero. |
| C Collisions between molecules and the walls are elastic. | D The kinetic energy of a given molecule is constant. |

2. The volume of an ideal gas in a container is increased at constant temperature. What happens as a result of this change?

- | | |
|---|---|
| A The frequency of collisions of the molecules with the container wall decreases. | B The intermolecular force between the molecules decreases. |
| C The speed of the particles increases. | D The volume of a gas particle increases. |

3. The temperature of a closed, rigid gas container is increased. What happens to the pressure and to the average speed of the particles?

- | | |
|--|--|
| A pressure increases; particles faster | B pressure decreases; particles faster |
| C pressure increases; no change in speed | D pressure decreases; no change in speed |

4. Which of the following are assumptions of the kinetic model of ideal gases?

- I. The particles undergo perfectly elastic collisions.
- II. The interactions between the particles are negligible when they are not colliding.
- III. The particles are in continuous random motion.

- A I and II only B I and III only C II and III only D I, II and III

5. Why does the pressure of an ideal gas increase when the temperature is increased at constant volume?

- | | |
|---|---|
| A The particles collide with each other more frequently. | B The time of contact between the molecules and the wall is reduced. |
| C The force of attraction between the molecules is reduced. | D The particles collide with the walls of the container more frequently and more energetically. |

6. Real gases deviate from ideal gas behaviour because their particles:

- | | |
|--------------------------|--|
| A have negligible volume | B have forces of attraction between them |
| C are polyatomic | D are not attracted to one another |

7. Which gas most closely approaches ideal behaviour at room temperature and pressure?

- A carbon dioxide B chlorine C hydrogen fluoride D neon

8. Identify the conditions under which carbon dioxide is best described by the ideal gas model.

- A low density and low pressure B low density and high pressure
C high density and low pressure D high density and high pressure

9. Why does nitrogen deviate from the ideal gas law at very high pressures?

- A Intermolecular forces increase the volume compared with that of an ideal gas. B Increasing the pressure decreases the volume of the molecules.
C Nitrogen molecules have a finite volume. D Collisions between molecules become inelastic at high pressure.

10. Identify the phenomenon that is best explained by the molecular kinetic model.

- A The difference between the Celsius and Kelvin temperature scales B The value of the universal gas constant
C The ideal gas equation D The existence of the noble gases

11. **HL Stretch** For a sample of a real gas, PV/nRT is found to be 0.87. What does this indicate?

- A The volume of the particles is significant. B Attractive intermolecular forces are significant.
C The gas is behaving ideally. D The collisions are inelastic.

12. Which change would cause the greatest deviation from ideal gas behaviour?

- A increasing the temperature at constant pressure B decreasing the pressure at constant temperature
C increasing the pressure at constant temperature D increasing the volume at constant temperature

13. Helium and argon are held at the same temperature. Which quantity is the same for both gases?

- A average speed of the particles B average kinetic energy of the particles
C density D rate of diffusion

14. Which pair of gases would be expected to deviate *least* from ideal behaviour under the same conditions?

- A H_2O and NH_3 B He and H_2 C Cl_2 and HCl D SO_2 and CO_2

15. Which statement about an ideal gas is correct?

- A It can be liquefied if the temperature is low enough. B Its particles have negligible volume.
C It obeys $PV = nRT$ only at STP. D All of its particles move at the same speed.

16. The temperature of a gas in kelvin is a measure of the:

- A total kinetic energy of the particles B average kinetic energy of the particles
C average speed of the particles D pressure exerted by the particles

17. For a fixed mass of an ideal gas at constant temperature, the graph of PV against P is:

- A a straight line through the origin with positive gradient B a curve that decreases towards zero
C a horizontal straight line D a straight line with a negative gradient

18. Which statement is *not* an assumption of the ideal gas model?

- A The particles are in continuous random motion. B All the particles have the same kinetic energy.
C The volume of the particles is negligible. D Collisions between particles are elastic.

19. **HL Stretch** At moderately high pressure, the measured pressure of a real gas is lower than that predicted by $PV = nRT$. The best explanation is that:

- A the particles occupy a significant volume B attractive forces reduce the energy of collisions with the wall
C the collisions with the wall are inelastic D the average kinetic energy has decreased

20. A real gas at constant pressure is cooled towards its boiling point. The deviation from ideal behaviour:

- A decreases B remains unchanged C increases D first increases, then decreases to zero

Part G — Practice: Paper 1B data-based

2 QUESTIONS · SL + HL

Question G1 — Compression factors of three gases

SL + HL

HL Stretch (d)-(f)

[12 marks]

The quantity PV/nRT is a direct measure of how far a gas departs from ideal behaviour. The table gives typical values for three gases, all measured at 300 K.

Pressure / 10^6 Pa	1	10	20	40	80
helium, He	1.00	1.05	1.10	1.20	1.40
nitrogen, N ₂	1.00	0.98	1.00	1.06	1.32
carbon dioxide, CO ₂	0.96	0.52	0.63	0.83	1.25

- (a) State the value of PV/nRT for an ideal gas, and explain why this value is independent of pressure. [2]
- (b) Describe how the value for carbon dioxide changes as the pressure is raised from 1×10^6 Pa to 8×10^7 Pa. [2]
- (c) Explain, in terms of the assumptions of the ideal gas model, why the carbon dioxide values first fall below 1 and then rise above it. [3]
- (d) Identify the gas whose behaviour is closest to ideal over the whole range and justify your choice. [2]
- (e) Explain why the value for helium never falls below 1. [2]
- (f) Predict, with a reason, how the carbon dioxide values would change if all the measurements were repeated at 600 K. [1]

Question G2 — Pressure and temperature of a fixed mass of gas

SL + HL

[13 marks]

A student measured the pressure of a fixed mass of air in a sealed rigid vessel of constant volume as the temperature was raised in a water bath. The results are shown below.

Temperature, $\theta / ^\circ\text{C}$	10	20	30	40	50	60	70	80
Pressure, P / kPa	99.0	102.5	105.8	109.2	112.5	116.0	119.3	122.7

The line of best fit through the data has the equation $P = 0.338\theta + 95.7$ (with P in kPa and θ in $^\circ\text{C}$).

- (a) State the independent variable, the dependent variable and one variable that must be controlled. [2]
- (b) Explain, using the ideal gas model, why the pressure increases as the temperature increases at constant volume. [3]
- (c) Use the equation of the line to determine the Celsius temperature at which the pressure of the gas would fall to zero. [2]
- (d) The accepted value of absolute zero is -273°C . Calculate the percentage error in the student's value and comment on the type of error that is most likely responsible. [3]
- (e) Sketch, on labelled axes, the graph that would be obtained if the same results were plotted with temperature in kelvin. State one feature of your sketch that is different from the graph obtained using $^\circ\text{C}$. [2]
- (f) State the assumption of the ideal gas model that would be least valid if the experiment were repeated at very low temperatures. [1]

Part H — Practice: Paper 2 structured 3 QUESTIONS · SL + HL**Question H1 — The ideal gas model** SL + HL [10 marks]

- (a) State **three** assumptions made in the kinetic model of an ideal gas. [3]
- (b) Explain, in terms of particles, the origin of the pressure exerted by a gas on the walls of its container. [2]
- (c) A sealed syringe contains a fixed mass of an ideal gas. The plunger is pushed in so that the volume is halved, while the temperature is held constant.
- (i) State the effect on the pressure of the gas. [1]
- (ii) Explain your answer in terms of the ideal gas model. [2]
- (d) A student states: "when the gas is compressed, the intermolecular forces between the particles increase, and that is why the pressure rises." Evaluate this statement. [2]

Question H2 — Real gases SL + HL HL Stretch (c) [11 marks]

- (a) Identify the **two** assumptions of the ideal gas model that are not valid for real gases, and state the effect of each on the pressure of the gas compared with the ideal value. [4]
- (b) Explain why deviations from ideal behaviour are greatest at high pressure and at low temperature. [4]
- (c) Ammonia ($M = 17.0 \text{ g mol}^{-1}$, boiling point $-33 \text{ }^{\circ}\text{C}$) and methane ($M = 16.0 \text{ g mol}^{-1}$, boiling point $-162 \text{ }^{\circ}\text{C}$) are compared at the same temperature and pressure. Deduce which shows the greater deviation from ideal behaviour and explain your answer. [3]

Question H3 — Testing the model quantitatively HL Stretch 1.5.4 [12 marks]

A sample of 2.00 mol of a gas is contained in a rigid vessel of volume 4.00 dm^3 at 250 K . The measured pressure is $9.00 \times 10^5 \text{ Pa}$. ($R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$)

- (a) Calculate the pressure that would be exerted by an ideal gas under these conditions. [2]
- (b) Calculate the value of PV/nRT for the gas. [2]
- (c) Deduce, with a reason, which of the two failing assumptions of the ideal gas model dominates under these conditions. [2]
- (d) One mole of the same gas, treated as an ideal gas, would occupy $6.65 \times 10^{-5} \text{ m}^3$ at 400 K and $5.00 \times 10^7 \text{ Pa}$. The molecules of one mole of this substance themselves occupy approximately $3.0 \times 10^{-5} \text{ m}^3$. Comment on the validity of the ideal gas model under these conditions, supporting your answer with a calculation. [3]
- (e) Despite these limitations, the ideal gas equation continues to be used throughout chemistry. Outline **two** reasons why it remains a useful model. [2]
- (f) State the conditions under which the gas would behave most nearly ideally. [1]

Part I — Mark schemes

FULL SOLUTIONS

I1 Paper 1A multiple choice — answers and distractor analysis

Q	Ans	Why it is right, and why the others are wrong
1	C	Elastic collisions are assumption 4. A — the model says nothing about mass; particles certainly have mass. B — the particles do exert force on the container; that is what pressure <i>is</i> . D — an individual molecule's kinetic energy changes at every collision; only the <i>average</i> is fixed by T .
2	A	Greater distance between wall collisions = lower frequency = lower pressure. B — in an <i>ideal</i> gas the intermolecular forces are already zero and cannot decrease (the most popular wrong answer). C — speed depends only on T , which is constant. D — particles never change size.
3	A	Higher $T \Rightarrow$ greater average $E_k \Rightarrow$ faster particles = more frequent and more energetic wall collisions = higher pressure. C and D wrongly separate speed from temperature; B reverses the pressure change.
4	D	All three are assumptions of the model (elastic collisions, negligible interactions except on collision, continuous random motion). Statement I is acceptable because "perfectly elastic" refers to the collisions.
5	D	Wall collisions become more frequent <i>and</i> more energetic. A — collisions between particles do not create wall pressure. B — contact time is not part of the model's explanation. C — there are no attractive forces in an ideal gas to reduce.
6	B	Attractive forces between particles are one of the two causes of deviation. A and D state the ideal assumptions, not the deviations. C — being polyatomic is not in itself the cause; the cause is the forces and finite volume.
7	D	Neon: monatomic, small, non-polar, only very weak London forces, boiling point 27 K. A and B are much larger and more polarisable; C , hydrogen fluoride, hydrogen bonds strongly.
8	A	Low density and low pressure means the particles are far apart, so both particle volume and intermolecular forces are negligible. High pressure or high density brings the molecules close and increases deviation.
9	C	At very high pressure the molecules' own volume is a significant fraction of the container, so $PV/nRT > 1$. A — intermolecular forces reduce, not increase, the apparent volume. B — molecules are not compressible. D — collisions remain essentially elastic.
10	C	The kinetic model explains the gas laws and hence the ideal gas equation. A is a matter of scale definition, B is an empirically measured constant, D is explained by electron configuration, not kinetic theory.

11	B	$PV/nRT < 1$ means the measured pressure (or volume) is below the ideal value, which is the signature of attractive forces. A would give a value greater than 1. C would give exactly 1. D is not a recognised cause of deviation.
12	C	Raising the pressure pushes the particles closer together, so both failing assumptions become worse. A and B and D all move the gas <i>towards</i> ideal behaviour.
13	B	Average kinetic energy depends only on absolute temperature. A and D — the lighter helium atoms move and diffuse faster at the same E_k . C — density depends on molar mass.
14	B	He and H ₂ have the smallest particles, fewest electrons, weakest London forces and the lowest boiling points (4 K and 20 K). A both hydrogen bond, C are large/polar, D SO ₂ is polar and both are large.
15	B	Negligible particle volume is assumption 1. A — an ideal gas has no attractions, so it could never condense. C — an ideal gas obeys $PV = nRT$ under <i>all</i> conditions, not just at STP. D — particles have a <i>range</i> of speeds.
16	B	Absolute temperature measures the <i>average</i> kinetic energy. A — the total also depends on how much gas is present. C — at the same T , different gases have different average speeds. D — pressure also depends on n and V .
17	C	At constant T , $PV = \text{constant}$, so plotting that constant against P gives a horizontal line. A is the graph of P against $1/V$; B is V against P .
18	B	The particles have a <i>range</i> of kinetic energies; only the average is fixed by the temperature. A , C and D are genuine assumptions.
19	B	Attractions slow the particles approaching the wall, so collisions are less energetic and the pressure is below the ideal value. A would raise the pressure above ideal. C is not a cause of deviation. D is false — the temperature has not changed.
20	C	Cooling towards the boiling point brings the gas closer to condensing: the particles move more slowly and the attractive forces have a proportionally greater effect, so deviation increases.

PATTERN TO NOTICE

Almost every wrong option in this set is one of four recurring traps: (i) claiming intermolecular forces change in an *ideal* gas; (ii) claiming particles change size; (iii) confusing *average* kinetic energy with the energy of an individual particle; (iv) confusing *average kinetic energy* (same for all gases at the same T) with *average speed* (different for each gas). Learn those four and you will not lose an MCQ mark on this topic.

I2 Paper 1B — Question G1 mark scheme [12]

(a) [2]

- $PV/nRT = 1$ for an ideal gas; (1)
- because an ideal gas obeys $PV = nRT$ exactly under all conditions — its particles have negligible volume and no intermolecular forces at any pressure, so nothing changes as P is raised. (1)

(b) [2]

- The value falls steeply from 0.96 to a minimum of about 0.52 at 1×10^7 Pa; (1)
- it then rises continuously, passing through 1 and reaching 1.25 at 8×10^7 Pa. (1)

Accept any answer that identifies both the initial decrease to a minimum and the subsequent increase above 1. Quoting figures from the table strengthens but is not required for the marks.

(c) [3]

- At moderate pressures the molecules are close enough for **attractive intermolecular forces** to act, slowing molecules approaching the wall, so the measured pressure is below the ideal value and the ratio falls below 1. (1)
- At very high pressures the molecules are forced very close together and their **own volume becomes a significant fraction** of the container volume. (1)
- Collisions with the wall then become more frequent than the model predicts, so the ratio rises above 1; the finite-volume effect has overtaken the attractive-force effect. (1)

(d) [2]

- **Nitrogen**; (1)
- its values stay closest to 1 over most of the range (0.98–1.06 up to 4×10^7 Pa), whereas helium is already at 1.20 and carbon dioxide has fallen to 0.52. (1)

Accept helium with full justification if the candidate argues from the smallest total spread only at the low-pressure end — but the data show helium departing steadily from 1 across the range, so nitrogen is the better answer.

(e) [2]

- Helium atoms are very small and non-polar with only extremely weak London dispersion forces (helium boils at 4 K), so attractive forces are negligible at 300 K. (1)
- With no significant attractions to lower the pressure, only the finite-volume effect operates, and that always raises the ratio above 1. (1)

(f) [1]

- The values would move closer to 1 (the minimum would be shallower), because at the higher temperature the average kinetic energy is greater relative to the intermolecular attractions, so the attractions have less effect. (1)

I3 Paper 1B — Question G2 mark scheme [13]

(a) [2]

- Independent variable: temperature. Dependent variable: pressure. (1)
- Controlled: the volume of the vessel (accept: the amount/mass of gas, i.e. no leaks). (1)

(b) [3]

- Increasing the temperature increases the average kinetic energy of the particles, so they move faster. (1)
- They therefore collide with the walls more frequently. (1)
- Each collision is also more energetic, so the force per unit area — the pressure — increases. (1)

(c) [2]

- $0 = 0.338\theta + 95.7 \Rightarrow \theta = -95.7 / 0.338$ (1)
- $\theta = -283\text{ °C}$ (1)

Accept -280 °C to -285 °C . Award both marks for a correct answer with no working.

(d) [3]

- percentage error = $|(-283) - (-273)| / 273 \times 100$ (1)
- $= 10 / 273 \times 100 = 3.7\%$ (1)
- The data are highly linear, so random error is small; a consistent offset of this kind points to a **systematic error** — for example a pressure gauge that is not zeroed correctly, a slow leak, or a length of connecting tubing whose gas is not at the temperature of the water bath. (1)

ECF applies: award the second and third marks for correct use of an incorrect answer from (c).

(e) [2]

- Sketch: a straight line of positive gradient, on axes labelled pressure / kPa (y) and temperature / K (x). (1)
- Difference: the line now passes through (or extrapolates to) the **origin**, showing that P is directly *proportional* to $T(K)$, whereas the Celsius graph has a large positive intercept and shows only a linear, not a proportional, relationship. (1)

(f) [1]

- The assumption that there are **no intermolecular forces** between the particles (assumption 2) — at low temperature the particles move slowly and the attractions are no longer negligible. (1)

I4 Paper 2 — Question H1 mark scheme [10]

(a) [3]

Any **three** of:

- the particles have negligible volume compared with the volume of the container; (1)
- there are no intermolecular forces between the particles; (1)
- the particles are in continuous random motion, with an average kinetic energy proportional to the absolute temperature; (1)
- all collisions are perfectly elastic (kinetic energy is conserved). (1)

(b) [2]

- The particles collide with the walls of the container; each collision exerts a force on the wall. (1)
- Pressure is the total force per unit area produced by the very large number of such collisions per second; because the motion is random it is the same on every wall. (1)

(c)(i) [1]

- The pressure **doubles** (accept: increases; $P \propto 1/V$). (1)

(c)(ii) [2]

- In the smaller volume the particles travel a shorter distance between collisions with the walls, so wall collisions become more frequent. (1)
- The temperature is unchanged, so the average speed and the energy of each collision are unchanged; the increase in frequency alone raises the force per unit area. (1)

(d) [2]

- The statement is **incorrect** for an ideal gas: the model assumes there are **no** intermolecular forces at all, so they cannot increase. (1)
- The correct explanation is the increased frequency of collisions with the walls. (Credit also: for a *real* gas the attractions do become significant on compression, but their effect is to *lower* the pressure below the ideal value, which is the opposite of what the student claims.) (1)

I5 Paper 2 — Question H2 mark scheme [11]

(a) [4]

- Assumption that the **particles have negligible volume** is invalid — real particles occupy a finite volume; (1)
- so the particles travel a shorter distance between wall collisions, collisions are more frequent, and the pressure is **higher** than the ideal value; (1)
- assumption that there are **no intermolecular forces** is invalid — real particles attract one another; (1)
- so particles approaching the wall are slowed, collisions are less energetic, and the pressure is **lower** than the ideal value. (1)

(b) [4]

- **High pressure:** the gas is compressed into a small volume so the particles are close together; (1)
- the particles' own volume is then a significant fraction of the container volume and the intermolecular forces act strongly, so both assumptions fail; (1)
- **Low temperature:** the particles have a lower average kinetic energy and move more slowly; (1)
- the attractive forces are therefore no longer negligible in comparison with the kinetic energy and have more time to act, so they significantly reduce the pressure. (1)

Award the marks for reasoning, not for restating the conditions. A candidate who writes only "because they are closer together and moving slower" scores 2 of 4.

(c) [3]

- **Ammonia** shows the greater deviation. (1)
- NH_3 is polar and forms **hydrogen bonds**, whereas CH_4 is non-polar with only weak **London dispersion** forces; (1)
- the much stronger attractions in ammonia make assumption 2 a far poorer approximation, so its pressure falls further below the ideal value. The far higher boiling point of ammonia (-33°C against -162°C) is direct evidence of these stronger forces, and the near-identical molar masses rule out particle size as the explanation. (1)

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I6 Paper 2 — Question H3 mark scheme [12]

(a) [2]

- $V = 4.00 \text{ dm}^3 = 4.00 \times 10^{-3} \text{ m}^3$; $P = nRT/V = (2.00 \times 8.31 \times 250) / (4.00 \times 10^{-3})$ (1)
- $P = 1.04 \times 10^6 \text{ Pa}$ ($1.0388 \times 10^6 \text{ Pa}$) (1)

Award only 1 mark if the volume is left in dm^3 (giving 1.04×10^3) but the method is otherwise correct.

(b) [2]

- $PV/nRT = (9.00 \times 10^5 \times 4.00 \times 10^{-3}) / (2.00 \times 8.31 \times 250) = 3600 / 4155$ (1)
- = **0.866** (1)

Accept the equivalent route $9.00 \times 10^5 / 1.0388 \times 10^6 = 0.866$ using the answer to (a); ECF applies.

(c) [2]

- The ratio is **less than 1**, so the measured pressure is below the ideal value. (1)
- The failure of the "no intermolecular forces" assumption therefore dominates: attractions pull the particles back from the walls, reducing the energy of wall collisions. (The finite-volume effect, which would raise the ratio above 1, is smaller at this pressure.) (1)

(d) [3]

- Fraction of the volume occupied by the molecules = $3.0 \times 10^{-5} / 6.65 \times 10^{-5}$ (1)
- = 0.45, i.e. about **45 %** (1)
- This is very far from negligible, so assumption 1 fails badly and the ideal gas equation is not valid under these conditions; the calculated volume would be seriously in error. (1)

(e) [2]

Any **two** of:

- it is sufficiently accurate (errors of only a few per cent) for most gases under ordinary laboratory conditions of temperature and pressure; (1)
- it is a single equation that applies to *every* gas, requiring no substance-specific constants (unlike van der Waals' equation, which needs *a* and *b* for each gas); (1)
- it is mathematically simple and gives a correct limiting behaviour — all real gases approach it as the pressure approaches zero; (1)
- it provides a clear conceptual picture linking pressure, volume, temperature and amount, and a baseline against which real deviations can be measured and understood. (1)

(f) [1]

- **High temperature and low pressure** (accept: low density). (1)

Where this topic goes next

STUDY PLAN

Next statement	What it adds
Structure 1.5.3 Molar volume	The molar volume of an ideal gas is constant at a given T and P : $22.7 \text{ dm}^3 \text{ mol}^{-1}$ at STP (273 K, 100 kPa). Investigating P - V - T relationships and analysing the graphs.
Structure 1.5.4 $PV = nRT$	Full quantitative work: solving for any variable, the combined gas law, and determining molar mass from experimental gas data.
Structure 2.2 Intermolecular forces	The reason <i>why</i> gases differ in their deviation — London dispersion, dipole-dipole and hydrogen bonding. Come back to section B3 once you have covered it.
Structure 1.1 Kinetic molecular theory	The parent model: states of matter, changes of state and the kelvin scale. Worth a re-read if the assumptions in Part A felt unfamiliar.

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