MEGA LECTURE

IB DP CHEMISTRY

Structure 1: Models of the Particulate Nature of Matter

Structure 1.5 Ideal Gases

Revision Notes · Standard and Higher Level

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

What the syllabus requires

Structure 1.5 models a gas as a swarm of independent particles in constant random motion, and uses that picture to link the pressure, volume, temperature and amount of a gas. Use this checklist to confirm you can do each item before an exam.

Understanding	You should be able to...
Kinetic molecular theory	State the assumptions of an ideal gas and use them to explain gas behaviour.
Gas variables and units	Convert temperatures to kelvin and handle volume and pressure units consistently.
The gas laws	State and apply Boyle's, Charles's and Gay-Lussac's laws and the combined gas law.
Ideal gas equation	Use PV = nRT to find any single variable, and rearrange it to find molar mass.
Avogadro's law	Relate gas volumes to mole ratios in reactions at the same temperature and pressure.
Real vs ideal gases	State the conditions under which real gases deviate from ideal behaviour, and explain why.

> **Exam note:** The whole of Structure 1.5 is common to SL and HL; there is no HL-only content in this topic. Points flagged "(HL)" in these notes are extension detail that deepens understanding and links to later HL topics such as Reactivity 2.2 (rates) and Structure 1.4 (the mole).

1. Kinetic molecular theory of an ideal gas

The kinetic molecular theory pictures a gas as a very large number of tiny particles in continuous, random, straight-line motion. An **ideal gas** is a simplified model gas that obeys the following assumptions exactly:

- The particles have **negligible volume** compared with the volume of the container: the gas is almost entirely empty space.
- There are **no intermolecular forces** between particles, either attractive or repulsive, except during collisions.
- Collisions between particles, and with the container walls, are perfectly **elastic** no kinetic energy is lost overall.
- The particles are in constant **random motion**, travelling in straight lines until they collide.
- The average kinetic energy of the particles is **proportional to the absolute (kelvin) temperature**: KE $\propto$ T. Doubling the kelvin temperature doubles the average kinetic energy.

How the model explains pressure

Gas pressure arises from the countless collisions of particles with the container walls. Each collision exerts a tiny force; the pressure is the total force of all collisions per unit area. Anything that makes the particles hit the walls harder or more often → raises the pressure. This single idea explains every gas law that follows.

> **Why "ideal"?** No real gas obeys these assumptions perfectly, because real particles do occupy space and do attract one another. The ideal model is nonetheless an excellent approximation for real gases at the ordinary conditions met in most exam questions.

2. The gas variables and their units

Four measurable quantities describe the state of a fixed sample of gas. Getting the units right is where most marks are won or lost.

Variable	Symbol	Common units	SI unit (use with PV = nRT)
Pressure	P	Pa, kPa, atm	pascal, Pa (= N m^{-2})
Volume	V	cm^3, dm^3, m^3	cubic metre, m^3
Temperature	T	°C, K	kelvin, K
Amount of substance	n	mol	mole, mol

Temperature: always convert to kelvin

Every gas calculation uses the **absolute (kelvin) temperature**. The kelvin scale starts at absolute zero (0 K = −273.15 °C), the temperature at which particle motion is at a minimum. Convert with:

T (K) = temperature in °C + 273

For example 25 °C = 298 K and −10 °C = 263 K. Using Celsius in any gas law is the single most common error in this topic.

Volume conversions

- 1 dm^3 = 1000 cm^3 = 1 litre.
- 1 m^3 = 1000 dm^3 = 1 × 10^6 cm^3.
- To convert cm^3 → m^3, divide by 1 × 10^6; dm^3 → m^3, divide by 1000.

Pressure conversions

- 1 kPa = 1000 Pa.
- 1 atm = 101.3 kPa = 101 300 Pa ≈ 1.01 × 10^5 Pa.
- Standard pressure in the IB data booklet is 100 kPa (1 × 10^5 Pa).

> **Standard temperature and pressure (STP):** In the current IB data booklet STP is 273 K (0 °C) and 100 kPa. Under these conditions one mole of any ideal gas occupies the **molar volume** V_m = 22.7 dm^3 mol^{-1} (2.27 × 10^{-2} m^3 mol^{-1}).

Worked example 1 — unit conversions

Convert (a) 45 °C to kelvin, (b) 350 cm^3 to m^3, (c) 2.5 atm to Pa.

(a) T = 45 + 273 = **318 K**.

(b) 350 cm^3 ÷ (1 × 10^6) = **3.5 × 10^{-4} m^3**.

(c) 2.5 × 101 300 = **2.5 × 10^5 Pa** (2.53 × 10^5 Pa).

3. The gas laws

Each gas law describes how two variables are related while the others are held constant. All follow directly from the kinetic theory of Section 1.

Boyle's law: pressure and volume

At constant temperature and amount, the pressure of a fixed mass of gas is **inversely proportional** to its volume:

$$P \propto 1/V \text{ so } P_1V_1 = P_2V_2$$

Squeeze a gas into half the volume and the particles hit the walls twice as often, so the pressure doubles. A graph of P against V is a curve (a hyperbola) falling away from both axes; a graph of P against 1/V is a straight line through the origin.

Worked example 2 — Boyle's law

A gas occupies 250 cm^3 at 100 kPa. What volume will it occupy at 400 kPa, the temperature being unchanged?

$P_1V_1 = P_2V_2 \rightarrow V_2 = P_1V_1 / P_2 = (100 \times 250) / 400$.

V_2 = 25 000 / 400 = **62.5 cm^3**. The pressure rose four-fold, so the volume fell to one quarter — a useful check.

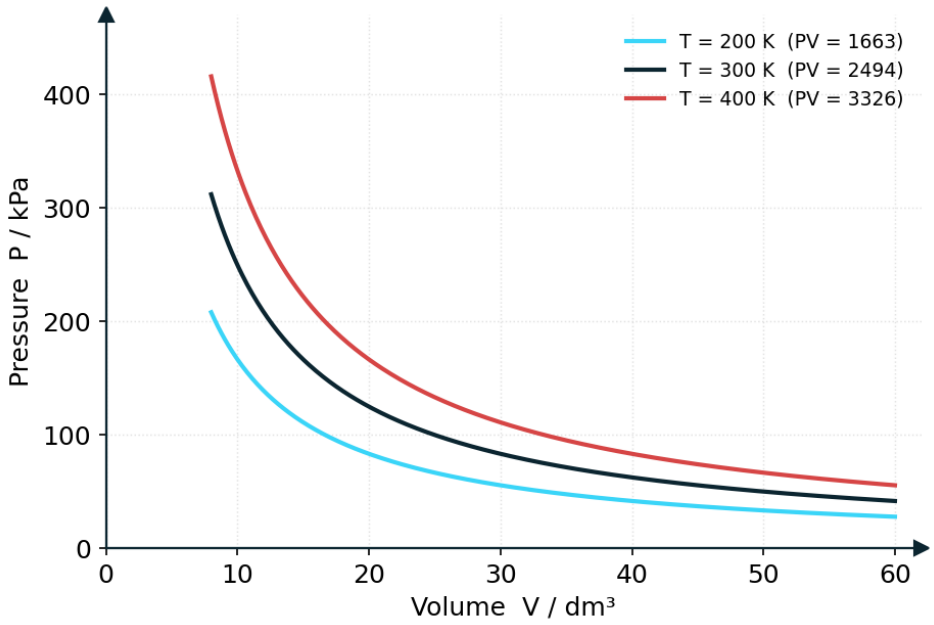

Figure 1. Boyle's law isotherms: at constant temperature $P \propto 1/V$, so the product $P \times V = nRT$ is constant along each curve (n = 1 mol; legend gives PV in kPa dm^3).

Charles's law: volume and temperature

At constant pressure and amount, the volume of a fixed mass of gas is **directly proportional** to its kelvin temperature:

$$V \propto T \text{ so } \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Heating a gas makes its particles move faster; to keep the pressure constant the gas must expand. A graph of V against T (in kelvin) is a straight line through the origin; extrapolated back, it meets the temperature axis at −273 °C (0 K), which is one way of locating absolute zero.

Worked example 3 — Charles's law

A balloon holds 2.0 dm^3 of gas at 300 K. To what volume does it expand when warmed to 450 K at constant pressure?

$V_1 / T_1 = V_2 / T_2 \rightarrow V_2 = V_1 \times T_2 / T_1 = 2.0 \times (450 / 300)$.

$V_2 = 2.0 \times 1.5 =$ **3.0 dm^3**. (Both temperatures already in kelvin — never use °C here.)

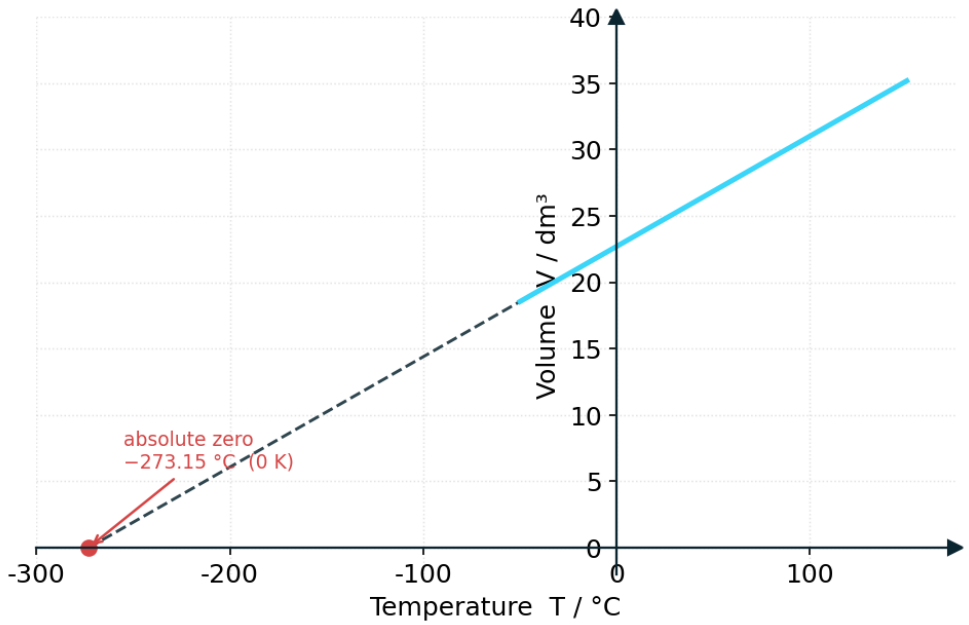

Figure 2. Charles's law at constant pressure: V ∝ T. Extrapolating the straight line to V = 0 gives an x-intercept at −273.15 °C, i.e. absolute zero (0 K).

Gay-Lussac's law: pressure and temperature

At constant volume and amount, the pressure of a fixed mass of gas is **directly proportional** to its kelvin temperature:

$$P \propto T \text{ so } P_1 / T_1 = P_2 / T_2$$

In a sealed rigid container the volume cannot change, so heating the gas makes the particles strike the walls faster and more often, raising the pressure. This is why aerosol cans carry a warning against heating. A graph of P against T (kelvin) is a straight line through the origin.

Worked example 4 — Gay-Lussac's law

A sealed rigid flask contains gas at 101 kPa and 27 °C. Find the pressure when it is heated to 127 °C.

Convert: $T_1 = 27 + 273 = 300$ K, $T_2 = 127 + 273 = 400$ K.

$P_2 = P_1 \times T_2 / T_1 = 101 \times (400 / 300) =$ **135 kPa**.

The combined gas law

When pressure, volume and temperature all change for a fixed amount of gas, the three simple laws combine into one relationship:

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

Each of Boyle's, Charles's and Gay-Lussac's laws is just this expression with one variable held constant. Temperatures must be in kelvin; pressure and volume units may be anything, provided the same unit is used on both sides.

Worked example 5 — combined gas law

A sample of gas occupies 500 cm^3 at 27 °C and 101 kPa. Find its volume at 127 °C and 202 kPa.

Convert temperatures: T_1 = 300 K, T_2 = 400 K.

Rearrange: $V_2 = P_1V_1T_2 / (T_1P_2)$ = (101 × 500 × 400) / (300 × 202).

V_2 = 20 200 000 / 60 600 = **333 cm^3**. The pressure doubled (halves V) while the temperature rose by a factor 4/3 (raises V), giving a net decrease.

Method for two-state problems: List P, V and T for state 1 and state 2, convert every temperature to kelvin, keep pressure and volume units matched across the equation, then rearrange for the single unknown. You never need R for these "before-and-after" questions.

Recognising the graphs

Exam questions often show a gas law as a graph and ask you to identify it or predict its shape. The table gathers the shapes you are expected to recognise; "origin" means the line, if extended, passes through (0, 0).

Law	Axes	Shape of graph
Boyle's law	P (y) against V (x)	Curve (hyperbola) falling away from both axes; never touches either axis
Boyle's law	P (y) against 1/V (x)	Straight line through the origin
Charles's law	V (y) against T in K (x)	Straight line through the origin
Charles's law	V (y) against T in °C (x)	Straight line cutting the T axis at −273 °C
Gay-Lussac's law	P (y) against T in K (x)	Straight line through the origin

Reading a straight line: A line through the origin signals direct proportionality (Charles's and Gay-Lussac's laws, in kelvin). If a volume-temperature or pressure-temperature line does **not** pass through the origin, the temperature axis is almost certainly in °C, not kelvin.

4. The ideal gas equation, PV = nRT

The gas laws combine with Avogadro's law into a single equation of state that links all four variables for any amount of gas:

$$PV = nRT$$

Here R is the **universal (ideal) gas constant**, R = 8.31 J K^{-1} mol^{-1} (given in the data booklet). Because R carries SI units, every other quantity **must** be in SI units when you use this equation:

Quantity	Symbol	Unit required with R = 8.31
Pressure	P	pascal, Pa

Quantity	Symbol	Unit required with R = 8.31
Volume	V	cubic metre, m^3
Amount	n	mole, mol
Temperature	T	kelvin, K

Finding any one variable

Rearrange PV = nRT for whatever is unknown: n = PV/RT, V = nRT/P, and so on. The two conversions that trip students up every time are pressure into Pa and volume into m^3.

Worked example 6 — finding the amount, n

Find the amount (in mol) of gas in a 2.00 dm^3 vessel at 98.0 kPa and 25.0 °C.

Convert: P = 98 000 Pa; V = 2.00 dm^3 = 2.00×10^{-3} m^3; T = 298 K.

n = PV / RT = $(98\,000 \times 2.00 \times 10^{-3}) / (8.31 \times 298)$.

n = 196 / 2476 = **0.0792 mol** (3 s.f.).

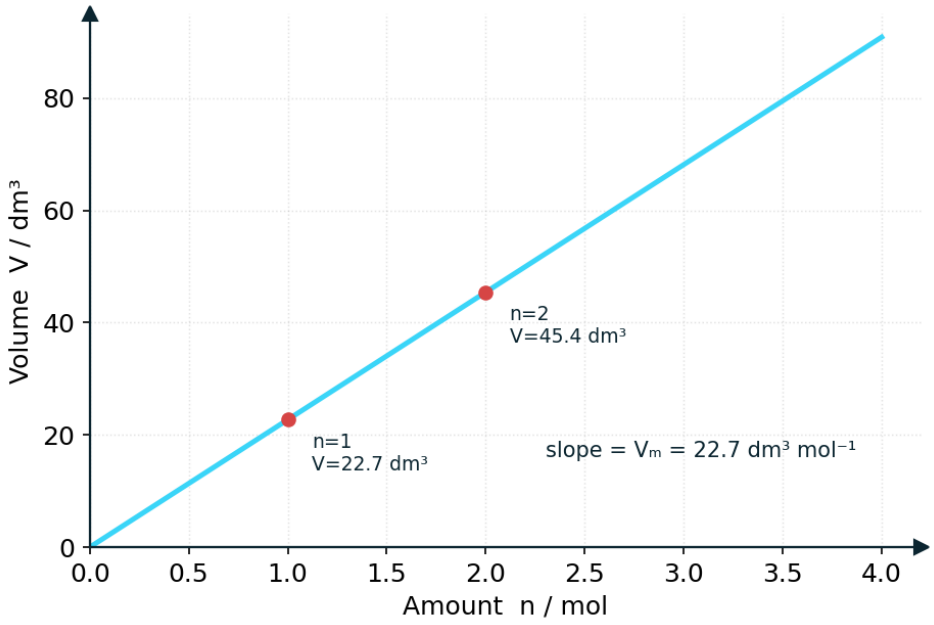

Figure 3. Ideal gas law PV = nRT: plotting V against n at STP (273.15 K, 100 kPa) gives a straight line of slope V/n = 22.7 dm^3 mol^{-1}; one mole of any ideal gas occupies 22.7 dm^3.

Finding molar mass

Because amount is mass divided by molar mass, n = m / M, substituting into PV = nRT and rearranging gives a direct route to the molar mass of a gas or volatile liquid:

$$M = mRT / (PV)$$

This is the basis of experiments that identify an unknown gas: measure the mass, volume, pressure and temperature of a sample, then compute M.

Worked example 7 — finding molar mass, M

0.500 g of a gas occupies 0.400 dm^3 at 100 kPa and 300 K. Find its molar mass.

Convert: P = 1.00 × 10^5 Pa; V = 0.400 dm^3 = 4.00 × 10^{-4} m^3.

M = mRT / (PV) = (0.500 × 8.31 × 300) / (1.00 × 10^5 × 4.00 × 10^{-4}).

M = 1246.5 / 40.0 = **31.2 g mol^{-1}**. (Numerator has units g × J mol^{-1}; the joules cancel with Pa m^3, leaving g mol^{-1}.)

Molar volume shortcut

At a fixed temperature and pressure every ideal gas has the same molar volume V_m = V/n. At STP (273 K, 100 kPa) V_m = 22.7 dm^3 mol^{-1}, so for a gas at STP you can use n = V / 22.7 (with V in dm^3) instead of the full equation — faster, and it avoids the SI conversions. Away from STP, use PV = nRT.

5. Avogadro's law and reacting gas volumes

Avogadro's law states that equal volumes of all gases, measured at the same temperature and pressure, contain equal numbers of particles (V ∝ n at fixed T and P). A direct and very useful consequence: for gases reacting at the same T and P, the ratio of reacting volumes equals the mole ratio in the balanced equation.

So the coefficients in an equation can be read straight off as volume ratios, without any need to work through moles. For example, in the synthesis of ammonia:

$$\mathbf{N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)}$$

1 volume of nitrogen reacts with 3 volumes of hydrogen to give 2 volumes of ammonia. So 50 cm^3 of N_2 reacts with 150 cm^3 of H_2 to form 100 cm^3 of NH_3, all measured at the same temperature and pressure.

Worked example 8 — volume of gas produced in a reaction

Excess dilute hydrochloric acid is added to 5.00 g of calcium carbonate. What volume of carbon dioxide is produced, measured at STP? ($M(CaCO_3)$ = 100.09 g mol^{-1}; V_m = 22.7 dm^3 mol^{-1} at STP.)

Equation: $CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$.

$n(CaCO_3)$ = 5.00 / 100.09 = 0.0500 mol, so $n(CO_2)$ = 0.0500 mol (1 : 1).

$V(CO_2) = n \times V_m$ = 0.0500 × 22.7 = **1.13 dm^3** (1130 cm^3).

6. Real gases versus the ideal model

Real gases obey PV = nRT closely under everyday conditions but deviate from it under extreme conditions. The deviations occur precisely because two of the ideal assumptions break down.

Ideal assumption	Why it fails in a real gas	When it matters most
Particles have negligible volume	Real particles occupy a real volume, which becomes significant once the particles are pushed close together.	High pressure (particles crowded)
No intermolecular forces	Real particles do attract one another; these attractions pull particles together and reduce the pressure they exert.	Low temperature (particles slow)

Conditions for deviation

A real gas behaves **least** ideally at **high pressure** and **low temperature**, and **most** ideally at **low pressure** and **high temperature**:

- **High pressure** forces the particles close together, so their own volume is no longer negligible next to the shrinking container volume, and attractions between them become important.
- **Low temperature** means the particles move slowly, so the weak intermolecular attractions have time to act and pull particles together, lowering the measured pressure below the ideal prediction.
- **Low pressure and high temperature** keep particles far apart and fast-moving, so both their volume and their attractions are negligible → behaviour is close to ideal.

Which gases behave most ideally?

Gases whose particles are **small** and have **weak intermolecular forces** behave most ideally. Helium and hydrogen are the closest to ideal: they are tiny and have only very weak London (dispersion) forces. Gases with larger molecules or stronger intermolecular forces — such as ammonia (hydrogen bonding) or carbon dioxide — deviate more, because their attractions and molecular volumes are greater.

> **HL link:** (HL) The same intermolecular forces that cause deviations from ideal behaviour (Structure 2.2) also set the boiling points of substances: a gas that deviates strongly at low temperature is one whose particles attract each other strongly and therefore condenses readily.

7. Common pitfalls

- Using °C instead of kelvin — every gas law and PV = nRT needs absolute temperature. Add 273 first, always.
- Forgetting the SI conversions in PV = nRT: pressure into Pa (not kPa) and volume into m^3 (not cm^3 or dm^3).
- Mixing units across the combined gas law, for example cm^3 on one side and dm^3 on the other.
- Quoting the molar volume as 24.0 dm^3 mol^{-1} (the old room-temperature value). The current IB STP value is **22.7 dm^3 mol^{-1}** at 273 K and 100 kPa.
- Reading reacting-volume ratios from unbalanced equations, or comparing gas volumes measured at different temperatures or pressures.
- Applying volume ratios to species that are not gases — Avogadro's law relates gas volumes only, never solids or liquids.

8. Quick reference

Result	Statement
Boyle's law (const T, n)	$P_1V_1 = P_2V_2$
Charles's law (const P, n)	$V_1 / T_1 = V_2 / T_2$
Gay-Lussac's law (const V, n)	$P_1 / T_1 = P_2 / T_2$
Combined gas law (const n)	$P_1V_1 / T_1 = P_2V_2 / T_2$
Ideal gas equation	$PV = nRT$, with $R = 8.31\ J\ K^{-1}\ mol^{-1}$
Molar mass	$M = mRT / (PV)$
Temperature conversion	T (K) = (°C value) + 273
Molar volume at STP	$V_m = 22.7\ dm^3\ mol^{-1}$ (273 K, 100 kPa)
Avogadro's law	Equal volumes of gases at same T, P contain equal numbers of particles
Least ideal conditions	High pressure and low temperature

9. Test yourself

Attempt all eight without notes, then check against the worked answers. Take $R = 8.31\ J\ K^{-1}\ mol^{-1}$ and $V_m = 22.7\ dm^3\ mol^{-1}$ at STP where needed.

1. A gas occupies 750 cm^3 at 95.0 kPa. It is compressed at constant temperature to 300 cm^3. Find the new pressure.

2. 40.0 cm^3 of gas at 25 °C is heated at constant pressure to 100 °C. Find its new volume.

3. A sealed rigid cylinder contains gas at 101 kPa and 300 K. Find the pressure after heating to 360 K.

4. 1.50 dm^3 of gas at 300 K and 100 kPa is cooled and compressed to 250 K and 120 kPa. Find its new volume.

5. Find the amount (mol) of gas in a 5.00 dm^3 container at 250 kPa and 350 K.

6. 1.20 g of a gas occupies 0.900 dm^3 at 100 kPa and 300 K. Find its molar mass.

7. Excess hydrochloric acid reacts with 0.240 g of magnesium: $Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$. Find the volume of hydrogen at STP. ($M(Mg) = 24.31\ g\ mol^{-1}$.)

8. In the reaction $2\ CO(g) + O_2(g) \rightarrow 2\ CO_2(g)$, what volume of oxygen reacts with 60 cm^3 of carbon monoxide, and what volume of carbon dioxide forms, all at the same T and P?

Answers

1. Boyle: $P_2 = P_1V_1/V_2 = (95.0 \times 750)/300 =$ **238 kPa**.

2. Charles: $T_1 = 298$ K, $T_2 = 373$ K; $V_2 = 40.0 \times (373/298) =$ **50.1 cm^3**.

3. Gay-Lussac: $P_2 = 101 \times (360/300) =$ **121 kPa**.

4. Combined: $V_2 = P_1V_1T_2/(T_1P_2) = (100 \times 1.50 \times 250)/(300 \times 120) = 37\ 500/36\ 000 =$ **1.04 dm^3**.

5. $PV = nRT$: $P = 250\ 000$ Pa, $V = 5.00 \times 10^{-3}\ m^3$; $n = (250\ 000 \times 5.00 \times 10^{-3})/(8.31 \times 350) = 1250/2908 =$ **0.430 mol**.

6. $M = mRT/(PV) = (1.20 \times 8.31 \times 300)/(1.00 \times 10^{5} \times 9.00 \times 10^{-4}) = 2991.6/90.0 =$ **$33.2\ g\ mol^{-1}$**.

7. $n(Mg) = 0.240/24.31 = 9.87 \times 10^{-3}$ mol; $n(H_2)$ = same (1 : 1). $V = 9.87 \times 10^{-3} \times 22.7 =$ **$0.224\ dm^3$** ($224\ cm^3$).

8. Volume ratio = mole ratio 2 : 1 : 2. So $O_2 = 60 \times (1/2) =$ **$30\ cm^3$**; $CO_2 = 60 \times (2/2) =$ **$60\ cm^3$**.