

Ideal gases

International Baccalaureate · IB Diploma · Chemistry · SL

Scope and prerequisites

This handout teaches Structure 1.5, Ideal gases, for the course first assessed in 2025. All of it is SL and HL content.

Prerequisites: the amount of substance n in mol and molar mass from Structure 1.4, and rearranging an equation. You will use the gas constant $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$ and the molar volume $22.7 \text{ dm}^3 \text{ mol}^{-1}$ from the data booklet.

The ideal gas model

■ Why does a sealed can burst in a fire?

Heat a sealed can and the pressure inside rises until the can splits. Pump air into a bicycle tyre and it gets harder. Squeeze a syringe of air with the end blocked and it pushes back. One simple model explains all three.

■ Five assumptions

The **kinetic molecular theory** pictures a gas as a huge number of tiny particles in fast, random motion. An **ideal gas** is one that obeys this model exactly:

1. The particles move randomly in straight lines until they collide.
2. The volume of the particles themselves is **negligible** compared with the volume of the container.
3. There are no **intermolecular forces** between the particles.
4. All collisions are **elastic**: no kinetic energy is lost.
5. The average kinetic energy of the particles is proportional to the kelvin temperature.

Pressure is the force per unit area that the particles exert when they hit the walls of the container. More hits per second, or harder hits, mean a higher pressure.

■ Temperature must be in kelvin

The **kelvin scale** starts at **absolute zero**, the lowest possible temperature, where the particles would have no kinetic energy.

$$T/\text{K} = t/^{\circ}\text{C} + 273$$

Kelvin temperature is proportional to the average kinetic energy of the particles. Celsius temperature is not: 40°C is not twice as hot as 20°C , but 400 K is twice 200 K .

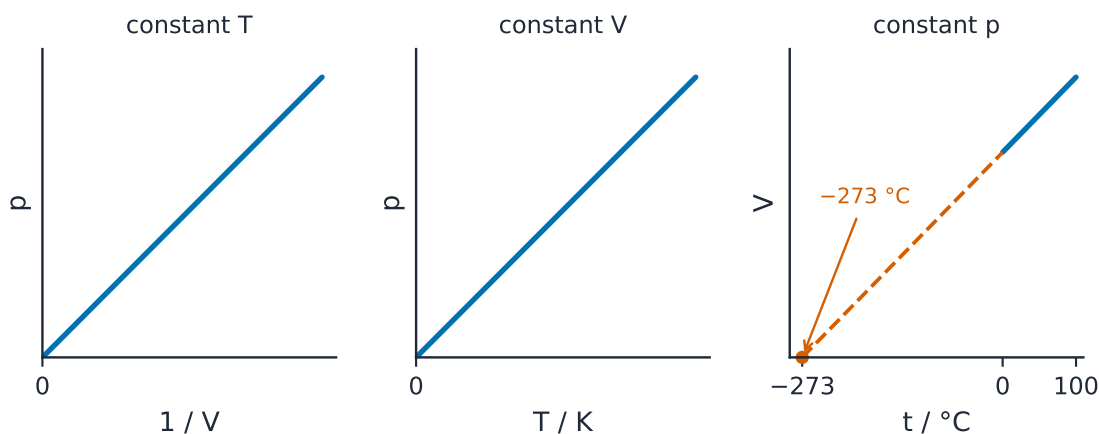
Every gas calculation uses kelvin.

How pressure, volume and temperature are linked

■ Three relationships for a fixed amount of gas

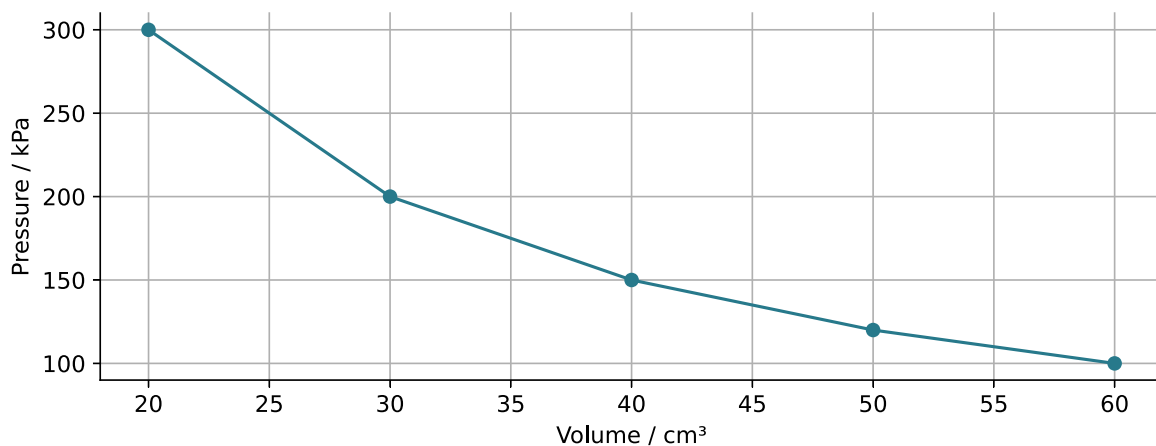
Change one quantity, hold one fixed, and watch the third. Each result has a particle explanation.

Held fixed	Relationship	Particle explanation
T	$p \propto \frac{1}{V}$, so pV is constant	smaller volume $\rightarrow$ particles hit the walls more often
V	$p \propto T$	hotter particles move faster, hit more often and harder
p	$V \propto T$	faster particles push the walls out until the pressure is back to its old value



Straight lines through the origin show proportionality; the volume line meets zero at -273°C , absolute zero

A plot of p against V is a curve, so it is hard to test. Plot p against $\frac{1}{V}$ instead: a straight line through the origin shows that $p \propto \frac{1}{V}$.



At constant temperature pV stays constant: $300 \times 20 = 100 \times 60 = 6000$

■ The combined gas law

The three relationships combine into one equation for a fixed amount of gas:

$$\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$$

Any units of pressure and volume work, as long as both sides use the same ones. Temperature must be in kelvin.

Worked example. A gas occupies 250 cm³ at 20 °C and 100 kPa. It is heated to 80 °C and the pressure rises to 150 kPa. Find the new volume.

1. Convert to kelvin: $T_1 = 293$ K, $T_2 = 353$ K.
2. Rearrange: $V_2 = \frac{p_1 V_1 T_2}{p_2 T_1}$.
3. Substitute: $V_2 = \frac{100 \times 250 \times 353}{150 \times 293} = 201$ cm³.
4. Check: the higher pressure squeezes the gas and the higher temperature expands it. The two effects nearly cancel, so a volume close to 250 cm³ is sensible.

Molar volume and the ideal gas equation

■ Equal volumes, equal numbers of particles

In an ideal gas the particles have negligible volume, so the volume of a gas depends only on **how many** particles there are, not on what they are. **Avogadro's law**

: equal volumes of all gases at the same temperature and pressure contain the same number of particles.

So one mole of any ideal gas has the same volume at a given temperature and pressure. This is the **molar volume**. At standard temperature and pressure (STP: 273 K and 100 kPa):

$$V_m = 22.7 \text{ dm}^3 \text{ mol}^{-1}$$

$$n = \frac{V}{V_m}$$

For example, 0.500 mol of oxygen or of carbon dioxide at STP each occupy $0.500 \times 22.7 = 11.4 \text{ dm}^3$.

Because volume is proportional to amount, the coefficients in a balanced equation for gases give their **volume ratio** too. In $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$, 10 cm^3 of nitrogen reacts with 30 cm^3 of hydrogen to give 20 cm^3 of ammonia, all measured at the same temperature and pressure.

■ The ideal gas equation

Putting amount into the combined law gives the **ideal gas equation**

:

$$pV = nRT$$

R is the **gas constant**, $8.31 \text{ J K}^{-1} \text{ mol}^{-1}$. The units must fit R :

Quantity	SI unit	Convert from
p	Pa	$1 \text{ kPa} = 10^3 \text{ Pa}$
V	m^3	$1 \text{ dm}^3 = 10^{-3} \text{ m}^3$; $1 \text{ cm}^3 = 10^{-6} \text{ m}^3$
T	K	add 273 to $^{\circ}\text{C}$

A useful shortcut: kPa multiplied by dm^3 also gives joules, so you may use kPa with dm^3 in the same equation.

Worked example: finding a molar mass. A 0.320 g sample of a gas occupies 180 cm^3 at 300 K and 100 kPa. Find its molar mass.

1. Convert: $p = 1.00 \times 10^5 \text{ Pa}$, $V = 180 \times 10^{-6} = 1.80 \times 10^{-4} \text{ m}^3$.
2. Amount: $n = \frac{pV}{RT} = \frac{1.00 \times 10^5 \times 1.80 \times 10^{-4}}{8.31 \times 300} = 7.22 \times 10^{-3} \text{ mol}$.
3. Molar mass: $M = \frac{m}{n} = \frac{0.320}{7.22 \times 10^{-3}} = 44.3 \text{ g mol}^{-1}$.
4. Interpret: this is close to 44.0 g mol^{-1} , so the gas could be carbon dioxide or propane.

Real gases

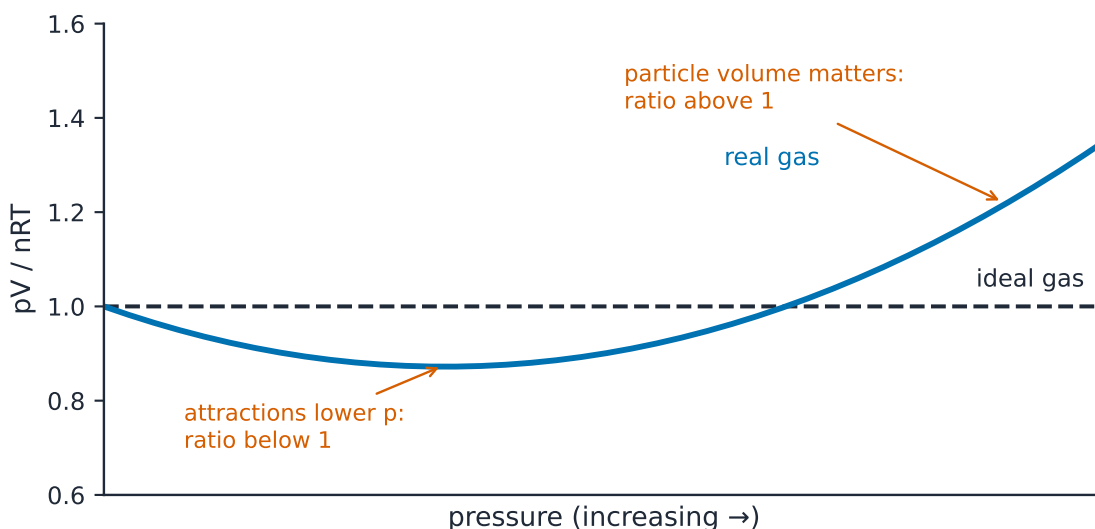
■ When the model breaks down

A **real gas** behaves almost ideally at high temperature and low pressure. It deviates most at **low temperature and high pressure**, because two assumptions stop being true.

- **High pressure.** The particles are pushed close together, so their own volume is no longer negligible. The space they can move in is smaller than the container volume, and the measured volume is **larger** than $pV = nRT$ predicts.
- **Low temperature.** The particles move slowly, so intermolecular attractions have time to act. A particle about to hit the wall is pulled back by its neighbours, so it hits less hard and the pressure is **lower** than predicted.

Gases with strong intermolecular forces, such as ammonia and water vapour, deviate more than gases with weak forces, such as helium and hydrogen.

We can compare a real gas with the model by plotting $\frac{pV}{nRT}$ against pressure. For an ideal gas this ratio is exactly 1 at every pressure.



Below 1, attractions win; above 1, the particles' own volume wins

■ Measuring gas behaviour in the lab

To test $p \propto T$, warm a sealed flask of air in a water bath connected to a pressure sensor. Let the gas reach the temperature of the bath before each reading, and keep within the temperature and pressure limits your teacher sets. **Never heat a sealed vessel that is not designed for it.**

Practice

1. Convert $-23\text{ }^{\circ}\text{C}$ and $127\text{ }^{\circ}\text{C}$ to kelvin.
2. A gas at 300 K and 120 kPa in a rigid container is cooled to 250 K. Find the new pressure.
3. Calculate the volume of 0.0400 mol of a gas at STP, in cm^3 .
4. Calculate the volume, in dm^3 , of 0.500 mol of gas at 298 K and 101 kPa.
5. 1.00 g of a gas occupies 0.763 dm^3 at 273 K and 100 kPa. Find its molar mass and suggest the gas.
6. Explain, in terms of particles, why the pressure of a gas rises when it is heated at constant volume.
7. State two assumptions of the ideal gas model, and explain why each fails at high pressure or low temperature.

Answers.

1. 250 K and 400 K.
2. $p_2 = 120 \times \frac{250}{300} = 100\text{ kPa}$.
3. $0.0400 \times 22.7 = 0.908\text{ dm}^3 = 908\text{ cm}^3$.
4. $V = \frac{nRT}{p} = \frac{0.500 \times 8.31 \times 298}{101} = 12.3\text{ dm}^3$ (kPa with dm^3).
5. $n = \frac{0.763}{22.7} = 0.0336\text{ mol}$; $M = \frac{1.00}{0.0336} = 29.8\text{ g mol}^{-1}$, close to 30: ethane, C_2H_6 , or nitrogen monoxide, NO.
6. The particles gain kinetic energy and move faster. They hit the walls more often and with more force, so the force per unit area rises.
7. Particle volume is negligible: at high pressure the particles are close together and their own volume is a large share of the container. No intermolecular forces: at low temperature the particles move slowly and attractions pull them together, lowering the pressure.

Exam tips

- Convert every temperature to **kelvin** before you substitute.
- In $pV = nRT$, use Pa with m^3 , or kPa with dm^3 ; never mix kPa with m^3 .
- In the combined gas law, any unit of p and V works as long as both sides match.
- “Deviates at high pressure and low temperature” earns a mark only with the reason: particle volume and intermolecular attractions.
- The molar volume $22.7\text{ dm}^3\text{ mol}^{-1}$ applies at STP (273 K, 100 kPa) only.