

The Gas Laws

What are the gas laws and how do you use the ideal gas law?

●●● Intermediate

States of Matter & Gases

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Before you start: it helps to understand [The Mole Concept](#) and [Temperature Scales: Celsius, Fahrenheit and Kelvin](#).

What is it?

The **gas laws** describe how four quantities of a gas are related:

Quantity	Symbol	Common units
Pressure	P	atm, kPa, mmHg
Volume	V	L, mL
Temperature	T	K (kelvin) only
Amount	n	mol

They all combine into a single equation, the **ideal gas law**:

$$PV = nRT$$

where R is the **gas constant**: $R = 0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$ or $R = 8.314 \text{ L}\cdot\text{kPa}/(\text{mol}\cdot\text{K})$.

🔑 KEY IDEA

Gas temperatures must always be in **kelvin**: $T(\text{K}) = T(^{\circ}\text{C}) + 273.15$ (many courses use 273). Gas laws use kelvin because it starts from absolute zero, where the volume of an ideal gas would shrink to zero. If a temperature is given in $^{\circ}\text{F}$, convert it to $^{\circ}\text{C}$ first: $T_{\text{C}} = (5^{\circ}\text{C}/9^{\circ}\text{F})(T_{\text{F}} - 32^{\circ}\text{F})$ (see [Temperature Scales](#)).

Why does it matter?

- **Everyday life:** tyres lose pressure in winter, aerosol cans warn “do not heat”, and a sealed bag of crisps puffs up on a plane.
- **Calculations:** with $PV = nRT$ you can turn a gas volume into moles and then use [stoichiometry](#) for reactions that make or use gases.
- **Molar mass:** measuring the mass and volume of a gas sample lets you identify the gas.

How does it work?

1. The simple gas laws

Each law keeps two quantities fixed and links the other two:

Law	Fixed	Relationship	Equation
Boyle's law	n, T	P and V inversely proportional	$P_1 V_1 = P_2 V_2$
Charles's law	n, P	V proportional to T	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Gay-Lussac's law	n, V	P proportional to T	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
Avogadro's law	P, T	V proportional to n	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$

2. Why gases behave this way

Gas particles move quickly and randomly, and pressure comes from their **collisions with the container walls**.

- **Smaller volume** (Boyle): the particles hit the walls more often, so pressure rises.
- **Higher temperature** (Gay-Lussac): the particles move faster and hit the walls harder and more often, so pressure rises. If the container can expand (Charles), the volume grows instead.
- **More particles** (Avogadro): more collisions, so at fixed pressure the gas takes up more space.

3. The combined gas law

When n stays the same but P , V and T all change:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Any unit of pressure or volume works here, as long as **both sides use the same units**. Temperature must be in kelvin.

4. The ideal gas law

$PV = nRT$ works for a single state of a gas: give it any three of P , V , n and T and it finds the fourth. Choose R to match your units:

- P in **atm**, V in **L**: use $R = 0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$
- P in **kPa**, V in **L**: use $R = 8.314 \text{ L}\cdot\text{kPa}/(\text{mol}\cdot\text{K})$

Molar volume: at 0 °C and 1 atm, one mole of any ideal gas occupies about **22.4 L**. At 25 °C and 1 atm it is about **24.5 L**.

∞ THINK OF IT LIKE THIS

Picture a crowd of people moving around a room. Shrink the room (smaller V) or make everyone run faster (higher T), and they bump into the walls more often: that's higher pressure. Add more people (more n) and, to keep the bumping the same, you need a bigger room.

🔍 MORE PRECISELY

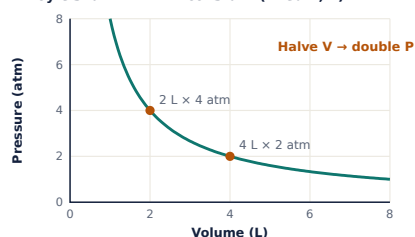
The ideal gas law assumes the particles have **no volume** and **no attractions** to each other. Real gases follow it closely at low pressure and high temperature, but deviate at high pressure and low temperature (especially near the point where they condense). “Standard temperature and pressure” (STP) is defined differently by different sources: 0 °C and 1 atm gives 22.4 L/mol, while the IUPAC definition of 0 °C and 1 bar (100 kPa) gives 22.7 L/mol. Check which one your course uses.

Visualise it

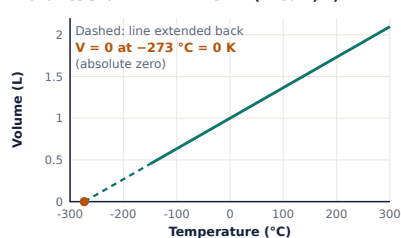
Two gas laws on graphs

Calculated for a fixed amount of an ideal gas.

Boyle's law: $P \times V = \text{constant}$ (fixed n , T)



Charles's law: $V \propto T$ in kelvin (fixed n , P)



Boyle's law is an inverse relationship; Charles's law is a straight line through absolute zero.

Worked example

WORKED EXAMPLE: Boyle's law

Question: A gas occupies 2.0 L at 1.0 atm. It is compressed to 0.50 L at constant temperature. What is the new pressure?

Step 1 $P_1 V_1 = P_2 V_2$, so:

$$\begin{aligned} P_2 &= \frac{P_1 V_1}{V_2} \\ &= \frac{(1.0 \text{ atm})(2.0 \text{ L})}{0.50 \text{ L}} \\ &= 4.0 \text{ atm} \end{aligned}$$

Step 2 The units of L cancel, leaving atm. $P_2 = \mathbf{4.0 \text{ atm}}$: the volume became 4 times smaller, so the pressure became 4 times larger.

WORKED EXAMPLE: Charles's law

Question: A balloon holds 3.00 L of gas at 27 °C. What is its volume at 127 °C, at the same pressure?

Step 1 Convert to kelvin: $T_1 = 27 \text{ °C} + 273.15 = 300.15 \text{ K}$ and $T_2 = 127 \text{ °C} + 273.15 = 400.15 \text{ K}$

Step 2 Rearrange $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ and substitute:

$$\begin{aligned} V_2 &= V_1 \times \frac{T_2}{T_1} \\ &= 3.00 \text{ L} \times \frac{400.15 \text{ K}}{300.15 \text{ K}} \\ &= 4.00 \text{ L} \end{aligned}$$

Step 3 The units of K cancel, leaving L. $V_2 = \mathbf{4.00 \text{ L}}$

 WORKED EXAMPLE: The ideal gas law

Question: What volume does 0.500 mol of gas occupy at 25 °C and 1.00 atm?

Step 1 Convert to kelvin: $T = 25\text{ °C} + 273.15 = 298.15\text{ K}$

Step 2 Pressure is in atm, so use $R = 0.08206\text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$. Rearrange $PV = nRT$ and substitute:

$$\begin{aligned} V &= \frac{nRT}{P} \\ &= (0.500\text{ mol}) \\ &\quad \times \left(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) \\ &\quad \times (298.15\text{ K}) \\ &\quad \div (1.00\text{ atm}) \\ &= 12.2\text{ L} \end{aligned}$$

Step 3 The units of mol, K and atm cancel, leaving L. $V = \mathbf{12.2\text{ L}}$

Common mistake

 COMMON MISTAKE: Using degrees Celsius

Heating a gas from 100 °C to 200 °C does **not** double its volume. In kelvin it goes from 373 K to 473 K, so the volume increases by a factor of only **1.27**.

 COMMON MISTAKE: Mixing units with R

With $R = 0.08206\text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$, pressure must be in **atm** and volume in **L**. With $R = 8.314\text{ L}\cdot\text{kPa}/(\text{mol}\cdot\text{K})$, use **kPa** and **L**. Convert first: 1 atm = 101.325 kPa = 760 mmHg; 1 L = 1000 mL.

 COMMON MISTAKE: Flipping the relationship in Boyle's law

Pressure and volume are **inversely** proportional: when one goes up, the other goes down. Check that your answer makes sense.

NOTATION NOTE

- Pressure may be given in atm, kPa, mmHg, torr (1 torr = 1 mmHg) or bar (1 bar = 100 kPa).
- Subscripts 1 and 2 mean “before” and “after” a change.
- Some books write $V \propto T$ (“V is proportional to T”).

Remember this

REMEMBER THIS

- Always use kelvin: $K = ^\circ C + 273.15$.
- Boyle $P_1 V_1 = P_2 V_2$; Charles V/T constant; Gay-Lussac P/T constant; Avogadro V/n constant.
- Combined: $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$; ideal: $PV = nRT$.
- $R = 0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$ or $R = 8.314 \text{ L}\cdot\text{kPa}/(\text{mol}\cdot\text{K})$; about 22.4 L/mol at 0 °C and 1 atm.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

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