

Gas laws and particle motion

AP Chemistry · Lesson

Quick answer: The ideal gas law, $PV = nRT$, links pressure, volume, amount and absolute temperature of a gas. Kinetic molecular theory explains it: pressure comes from particle collisions, and temperature measures average kinetic energy. Partial pressures add up in mixtures (Dalton's law), and real gases deviate at high pressure and low temperature.

Level: AP Chemistry · **Curriculum:** AP Chem Unit 3

What does the ideal gas law say?

Definition

An **ideal gas** is a model in which particles have no volume and exert no attractive or repulsive forces on one another. Real gases behave almost ideally at high temperature and low pressure.

Formula

$$PV = nRT$$

P = pressure (kPa or atm), V = volume (L), n = amount of gas (mol), T = temperature (K),
 $R = 8.314 \text{ L}\cdot\text{kPa}/(\text{mol}\cdot\text{K}) = 0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$. Choose the R whose units match your pressure and volume.

The ideal gas law combines older laws. Doubling the pressure of a fixed sample at constant temperature halves its volume; heating at constant pressure expands it.

Law	Held constant	Relationship	Form
Boyle	n, T	P and V inversely proportional	$P_1V_1 = P_2V_2$
Charles	n, P	V proportional to T (kelvin)	$V_1/T_1 = V_2/T_2$
Gay-Lussac	n, V	P proportional to T (kelvin)	$P_1/T_1 = P_2/T_2$
Avogadro	P, T	V proportional to n	$V_1/n_1 = V_2/n_2$

Worked example 1

Diego in Texas fills a rigid 4.00 L cylinder with 0.250 mol of helium at 25 °C. Find the pressure.

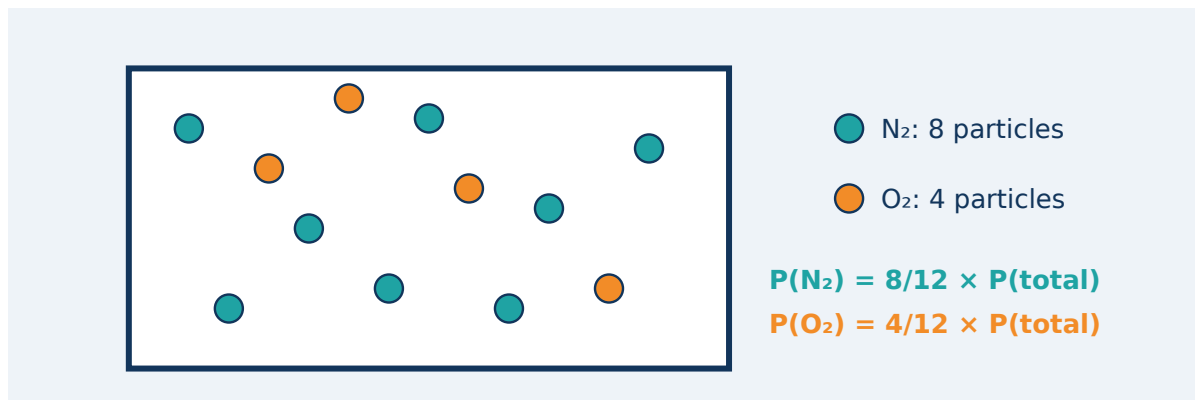
$$T = 25 + 273.15 = 298.15 \text{ K}$$

$$P = nRT/V = (0.250 \times 8.314 \times 298.15) / 4.00$$

$$\mathbf{P = 155 \text{ kPa}}$$

How do gases in a mixture share the pressure?

Gas particles are so far apart that each type behaves as if the others were absent. **Dalton's law** states that the total pressure of a mixture is the sum of the partial pressures. The partial pressure of gas A is $P_A = x_A \times P_{\text{total}}$, where the mole fraction $x_A = n_A/n_{\text{total}}$.



Each gas contributes to the total pressure in proportion to its share of the particles.

Worked example 2

A 10.0 L flask at 300 K holds 0.400 mol N_2 and 0.100 mol O_2 . Find the total pressure and the partial pressure of N_2 .

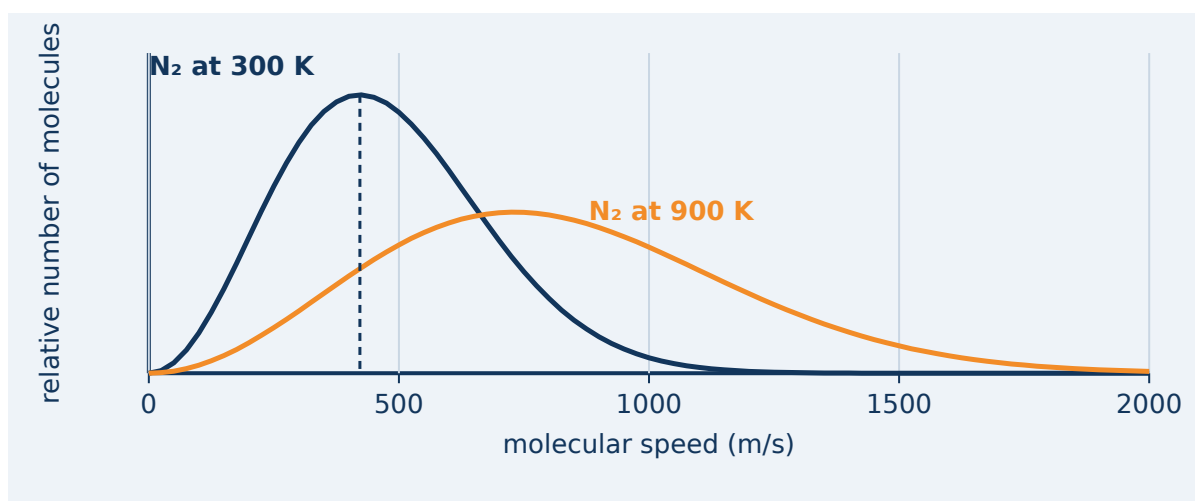
$$n_{\text{total}} = 0.500 \text{ mol, so } P_{\text{total}} = 0.500 \times 8.314 \times 300 / 10.0 = 124.7 \text{ kPa}$$

$$x(\text{N}_2) = 0.400/0.500 = 0.800, \text{ so } P(\text{N}_2) = 0.800 \times 124.7$$

$$P(\text{N}_2) = 99.8 \text{ kPa, } P_{\text{total}} = 125 \text{ kPa}$$

What does kinetic molecular theory explain?

The **kinetic molecular theory** (KMT) says that gas particles move randomly in straight lines, collide elastically with the walls and each other, and have an average kinetic energy proportional to the absolute temperature. Pressure arises from the sum of countless wall collisions. At the same temperature all gases have the same average kinetic energy, so lighter molecules move faster.



Speed distribution of N_2 . Raising the temperature shifts the peak right and spreads the curve; the area stays the same.

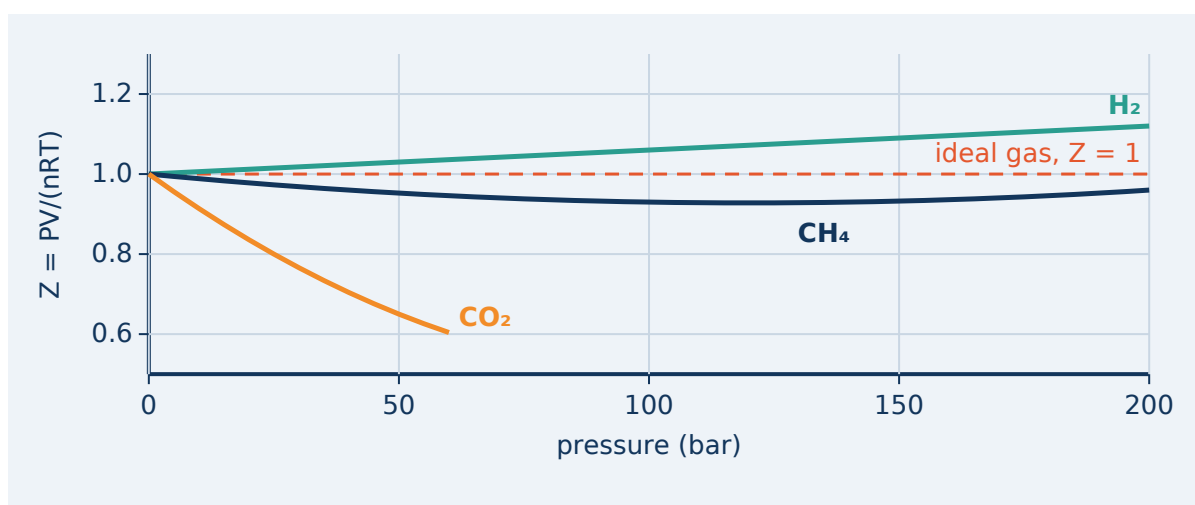
Key law

Average molar kinetic energy = $(3/2)RT$, and the root-mean-square speed is $u_{\text{rms}} = \sqrt{(3RT/M)}$, with M in kg/mol, T in K and $R = 8.314 \text{ J/(mol}\cdot\text{K)}$.

Gas	M (kg/mol)	u_{rms} at 300 K (m/s)
H ₂	0.002016	1930
He	0.004003	1370
N ₂	0.028014	517
CO ₂	0.04401	412

When do real gases deviate from ideal behaviour?

Two assumptions fail. At high pressure the particles take up a noticeable share of the volume, which makes the real volume larger than predicted. At low temperature attractions pull particles together and make the pressure smaller than predicted. The van der Waals equation corrects both: $(P + an^2/V^2)(V - nb) = nRT$, where a measures attractions and b the particle volume.



Schematic compressibility factors at 300 K. Z below 1 means attractions dominate; Z above 1 means particle volume dominates.

Common mistake

Never put the temperature in °C into $PV = nRT$. Always convert to kelvin. Also remember that gas collected over water contains water vapour, so subtract its partial pressure before using the gas law.

Lina's tip

Write the units next to every number first. If kPa and L go in, use $R = 8.314$; if atm and L go in, use 0.08206. When units cancel to leave mol, you are safe.

Key points to remember

- $PV = nRT$ with $R = 8.314 \text{ L}\cdot\text{kPa}/(\text{mol}\cdot\text{K})$ or $0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$; T must be in kelvin.
- Dalton: $P_{\text{total}} = \text{sum of partial pressures}$; $P_A = x_A \times P_{\text{total}}$.
- Root-mean-square speed $u_{\text{rms}} = \sqrt{3RT/M}$, M in kg/mol ; lighter gases move faster at the same T .
- Real gases deviate at high pressure (particle volume) and low temperature (attractions).
- $1 \text{ atm} = 101.3 \text{ kPa} = 760 \text{ mmHg}$.

Gas laws and particle motion: frequently asked questions

Why must gas temperature be in kelvin?

The gas laws are proportionalities to absolute temperature, which is zero when particle motion stops. A Celsius temperature of 0 does not mean zero kinetic energy, so substituting $^{\circ}\text{C}$ would give wrong results.

Which value of R should I use?

Use the one whose units match your data. With pressure in kPa and volume in litres use $8.314 \text{ L}\cdot\text{kPa}/(\text{mol}\cdot\text{K})$; in atm and litres use $0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$; for speeds and energies use $8.314 \text{ J}/(\text{mol}\cdot\text{K})$.

Why do lighter gases move faster at the same temperature?

At a given temperature all gases have the same average kinetic energy, $\frac{1}{2}mv^2$. With a smaller mass, the speed must be larger to give the same energy, which is why helium atoms are faster than nitrogen molecules.

What is gas collected over water?

A gas bubbled into an inverted container of water is saturated with water vapour. The total pressure equals the partial pressure of the gas plus the vapour pressure of water at that temperature, so subtract it first.

When does a real gas behave most ideally?

At high temperature and low pressure. Particles are then far apart and moving fast, so their own volume and the attractions between them are negligible compared with the volume and energy of the sample.