

Gas Mixtures and Gas Stoichiometry

How do you handle mixtures of gases and the volumes of gases in reactions?

●●○ Intermediate

States of Matter & Gases

Last reviewed 5 October 2026

Before you start: it helps to understand [The Gas Laws](#) and [Stoichiometry](#).

What is it?

In a mixture of gases that don't react, each gas behaves as if it were alone in the container. The pressure each gas would exert on its own is its **partial pressure**. **Dalton's law** says the total pressure is the sum of the partial pressures:

$$P_{\text{total}} = P_{\text{A}} + P_{\text{B}} + P_{\text{C}} + \dots$$

Because pressure depends only on the number of moles (at fixed V and T), each gas's share of the pressure equals its share of the moles, its **mole fraction** x :

$$P_{\text{A}} = x_{\text{A}} P_{\text{total}}$$
$$x_{\text{A}} = \frac{n_{\text{A}}}{n_{\text{total}}}$$

KEY IDEA

For gases, **volume is a measure of moles**. At the same temperature and pressure, equal volumes of gases contain equal numbers of moles (Avogadro's law). So $PV = nRT$ turns any gas volume into moles, and moles are the link to the balanced equation.

Why does it matter?

- **Breathing and diving.** Air is about 21 % oxygen, so the partial pressure of oxygen at sea level is about 0.21 atm. At high altitude it is lower, which is why climbers carry oxygen; divers breathing compressed air risk taking in too much nitrogen.
- **Collecting gases in the lab.** Gases collected over water are mixed with water vapour, which has to be subtracted.
- **Airbags, baking and industry.** Calculating the volume of gas a reaction produces is gas stoichiometry.

How does it work?

1. Partial pressures

Find each partial pressure with $P_A = n_A RT/V$, or from the mole fraction. Add them to get the total.

2. Collecting a gas over water

A gas bubbled into an upturned bottle of water becomes saturated with water vapour. The pressure inside equals atmospheric pressure, so:

$$P_{\text{gas}} = P_{\text{total}} - P_{\text{H}_2\text{O}}$$

The vapour pressure of water depends only on temperature (23.8 mmHg at 25 °C).

3. Molar volume

One mole of any ideal gas occupies **22.4 L at 0 °C and 1 atm**, and **24.5 L at 25 °C and 1 atm**. (Check which “standard” conditions your course uses; IUPAC STP, 0 °C and 1 bar, gives 22.7 L/mol.)

4. Gas stoichiometry

1. Convert what you know to **moles** (from mass, or from a gas volume with $n = PV/RT$).
2. Use the **mole ratio** from the balanced equation.
3. Convert to what you want (mass, or a gas volume with $V = nRT/P$).

THINK OF IT LIKE THIS

Partial pressures are like voices in a choir. Each singer adds to the total sound, and a section with more singers is louder. The total volume of sound is the sum of every section, and each section's share depends only on how many singers it has.

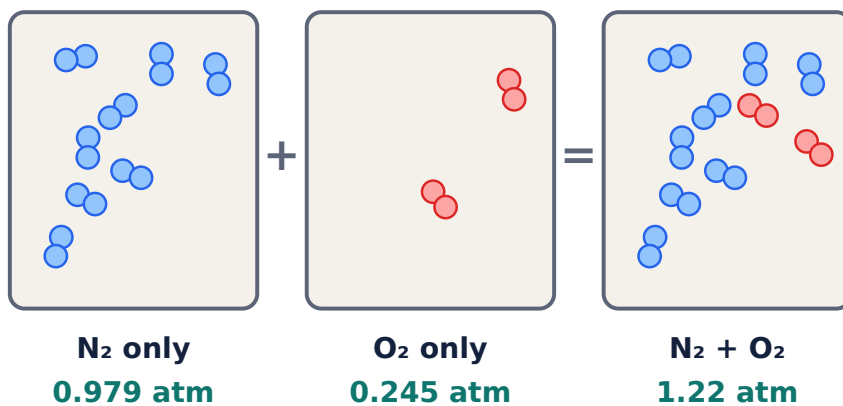
MORE PRECISELY

Dalton's law holds exactly for ideal gases and very closely for real gases at ordinary pressures. Because the partial pressure is proportional to moles, mole fraction is also the **volume fraction** of an ideal-gas mixture: air is about 78 % N₂, 21 % O₂ and 1 % Ar by volume and by moles. Vapour pressure tables give the partial pressure of water above liquid water at each temperature; at 100 °C it reaches 760 mmHg, which is why water boils there at 1 atm.

Visualise it

Dalton's law of partial pressures

Same 5.00 L flask, same temperature (25 °C)



0.200 mol N₂ and 0.0500 mol O₂: $P(\text{total}) = P(\text{N}_2) + P(\text{O}_2)$

Dalton's law: in the same container, each gas contributes its own partial pressure.

Worked example

WORKED EXAMPLE: Partial pressures

Question: A 5.00 L flask at 25 °C contains 0.200 mol of N₂ and 0.0500 mol of O₂. Find each partial pressure and the total pressure.

Step 1 $T = 25 + 273.15 = 298.15$ K. It saves work to find RT first:

$$\begin{aligned} RT &= 0.08206 \frac{\text{L atm}}{\text{mol K}} \\ &\times 298.15 \text{ K} \\ &= 24.47 \text{ L atm/mol} \end{aligned}$$

Step 2 Nitrogen:

$$\begin{aligned} nRT &= 0.200 \text{ mol} \\ &\times 24.47 \frac{\text{L atm}}{\text{mol}} \\ &= 4.894 \text{ L atm} \\ P_{\text{N}_2} &= \frac{nRT}{V} \\ &= \frac{4.894 \text{ L atm}}{5.00 \text{ L}} \\ &= 0.979 \text{ atm} \end{aligned}$$

Step 3 Oxygen has a quarter of the moles, so $P_{\text{O}_2} = 0.245$ atm.

Step 4 Total: $P = 0.979 \text{ atm} + 0.245 \text{ atm} = 1.22 \text{ atm}$

 WORKED EXAMPLE: A gas collected over water

Question: Hydrogen is collected over water at 25 °C. The volume is 0.250 L and the total pressure is 755 mmHg. How many moles of H₂ were collected?

Step 1 Pressure of the dry hydrogen:

$$\begin{aligned}P_{\text{H}_2} &= 755 \text{ mmHg} \\&\quad - 23.8 \text{ mmHg} \\&= 731.2 \text{ mmHg} \\&\quad \times \frac{1 \text{ atm}}{760 \text{ mmHg}} \\&= 0.9621 \text{ atm}\end{aligned}$$

Step 2 Moles, using $RT = 24.47 \text{ L atm/mol}$ at 25 °C (from the first example):

$$\begin{aligned}n &= \frac{PV}{RT} \\&= \frac{0.9621 \text{ atm} \times 0.250 \text{ L}}{24.47 \text{ L atm/mol}} \\&= 9.83 \times 10^{-3} \text{ mol}\end{aligned}$$

 WORKED EXAMPLE: Volume of gas from a reaction

Question: What volume of CO₂, at 25 °C and 1.00 atm, is released when 10.0 g of CaCO₃ decomposes? $\text{CaCO}_3 \longrightarrow \text{CaO} + \text{CO}_2$

Step 1 Moles of CaCO₃: $\frac{10.0 \text{ g}}{100.09 \text{ g/mol}} = 0.09991 \text{ mol}$, so 0.09991 mol CO₂ (1 : 1).

Step 2 Volume, with $RT = 24.47 \text{ L atm/mol}$ at 25 °C:

$$\begin{aligned}nRT &= 0.09991 \text{ mol} \\&\quad \times 24.47 \frac{\text{L atm}}{\text{mol}} \\&= 2.4448 \text{ L atm} \\V &= \frac{nRT}{P} = \frac{2.4448 \text{ L atm}}{1.00 \text{ atm}} \\&= 2.44 \text{ L}\end{aligned}$$

 WORKED EXAMPLE: Using molar volume

Question: 5.00 g of potassium chlorate is heated: $2 \text{KClO}_3 \rightarrow 2 \text{KCl} + 3 \text{O}_2$. What volume of O_2 forms at 0°C and 1 atm?

$$\begin{aligned} & 5.00 \text{ g } \text{KClO}_3 \\ & \times \frac{1 \text{ mol } \text{KClO}_3}{122.55 \text{ g } \text{KClO}_3} \\ & \times \frac{3 \text{ mol } \text{O}_2}{2 \text{ mol } \text{KClO}_3} \\ & \times \frac{22.4 \text{ L } \text{O}_2}{1 \text{ mol } \text{O}_2} \\ & = 1.37 \text{ L } \text{O}_2 \end{aligned}$$

Common mistake

 COMMON MISTAKE: Forgetting the water vapour

A gas collected over water is not pure. Subtract the vapour pressure of water before using $PV = nRT$, or the moles of gas will be too high.

 COMMON MISTAKE: Using 22.4 L/mol at room temperature

22.4 L/mol applies only at **0°C and 1 atm**. At 25°C and 1 atm, a mole of gas occupies 24.5 L. When conditions are not standard, use $PV = nRT$.

 COMMON MISTAKE: Mixing pressure units

With $R = 0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$, pressure must be in atm. Convert mmHg ($\div 760$) or kPa ($\div 101.325$) first.

 NOTATION NOTE

- x_A (or χ_A) is the mole fraction; mole fractions of all components add up to 1.
- 1 mmHg = 1 torr.

Remember this

REMEMBER THIS

- Dalton: $P_{\text{total}} = P_{\text{A}} + P_{\text{B}} + \cdots$; $P_{\text{A}} = x_{\text{A}}P_{\text{total}}$.
- Over water: $P_{\text{gas}} = P_{\text{total}} - P_{\text{H}_2\text{O}}$.
- Molar volume: 22.4 L/mol at 0 °C, 1 atm; 24.5 L/mol at 25 °C, 1 atm.
- Gas stoichiometry: gas volume $\leftrightarrow$ moles with $PV = nRT$; moles $\leftrightarrow$ moles with the equation.

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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