

Gases

CHEM 50 · Dr. Karmach

By the end of this unit, you can...

- State the conditions of STP and use the molar volume (22.4 L/mol) to convert between moles and volume of a gas
- Apply the combined gas law to relate the pressure, volume, and temperature of a gas sample
- Use the ideal gas law $PV = nRT$ to solve for any variable, a gas density, or a molar mass
- Apply Dalton's law of partial pressures, including a gas collected over water
- Carry mole ratios through gas-stoichiometry problems at STP and non-STP conditions
- Use Graham's law to compare effusion rates of gases from their molar masses

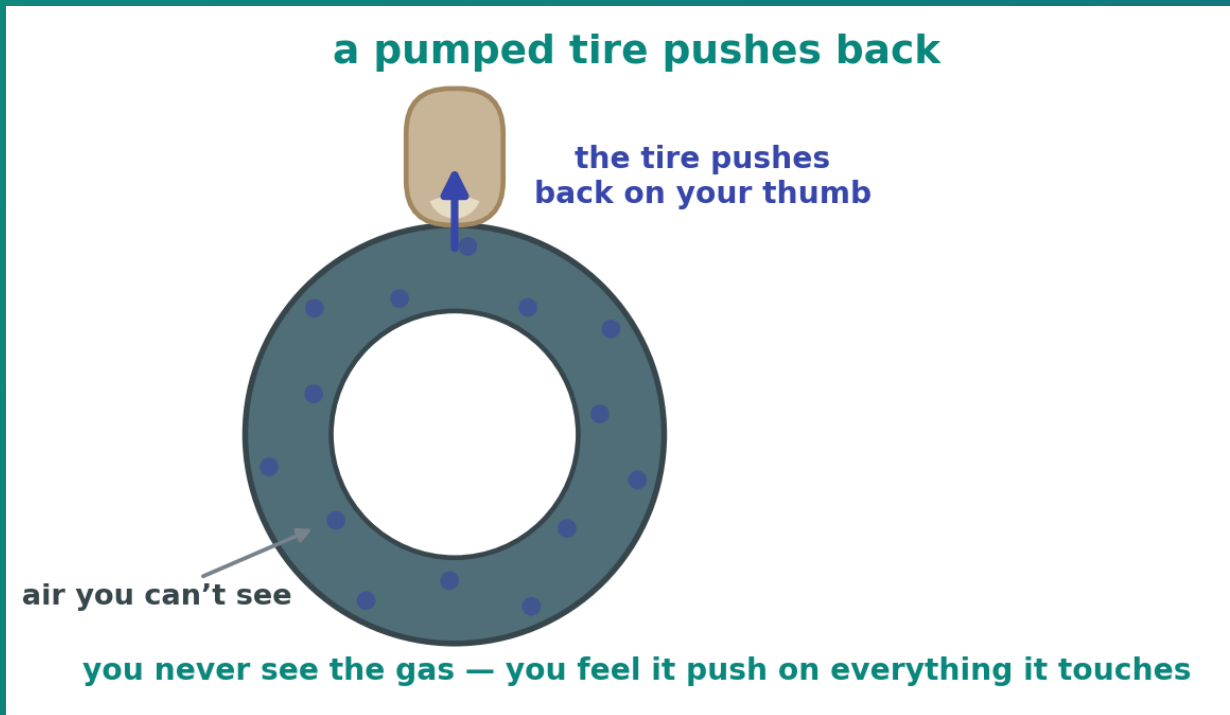
Today's route

1. **Gases & Pressure**
2. **Molar Volume at STP**
3. **The Combined Gas Law**
4. **The Ideal Gas Law**
5. **Dalton's Law of Partial Pressures**
6. **Collecting a Gas Over Water**
7. **Gas Stoichiometry**
8. **Kinetic-Molecular Theory**
9. **Real Gases and the van der Waals Equation**

1 • Gases & Pressure

Define gas pressure as the push of molecular collisions on a container's walls, and convert any pressure among atm, mmHg, torr, kPa, and psi.

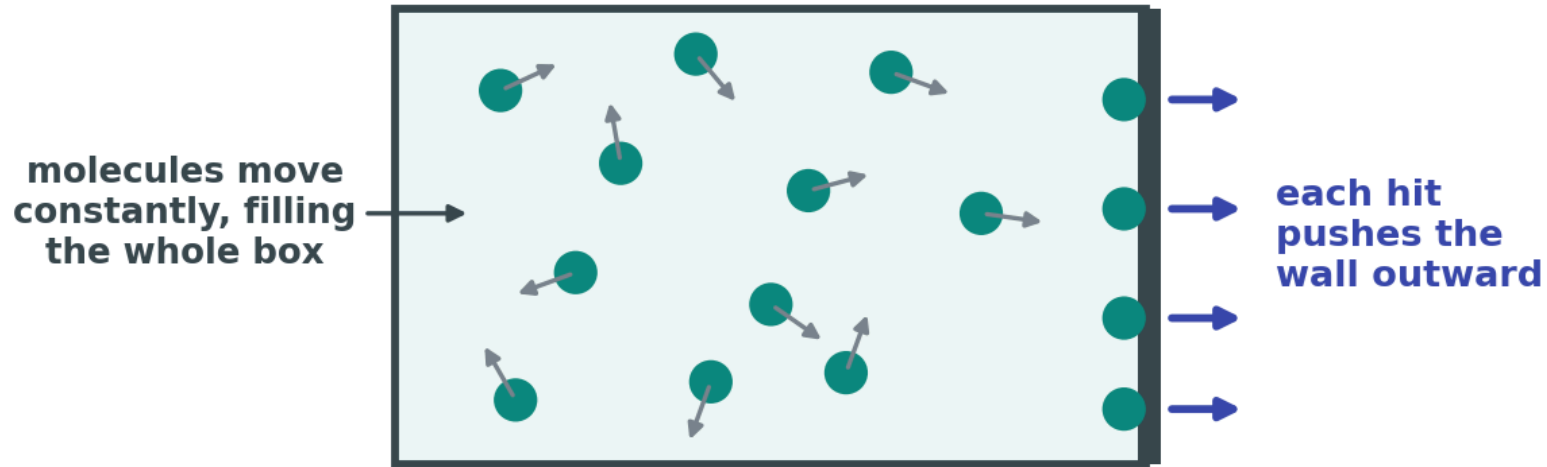
A bike tire pushes back



You can't see the air inside a bike tire. Press on the tread and it pushes right back, the same from every side.

A gas pushes on its container

a gas pushes on the walls it touches



countless collisions, spread over the wall's area — that push is the pressure

A gas has no shape of its own. Its molecules move constantly and fill the whole container, striking every wall. Each collision is a tiny push outward.

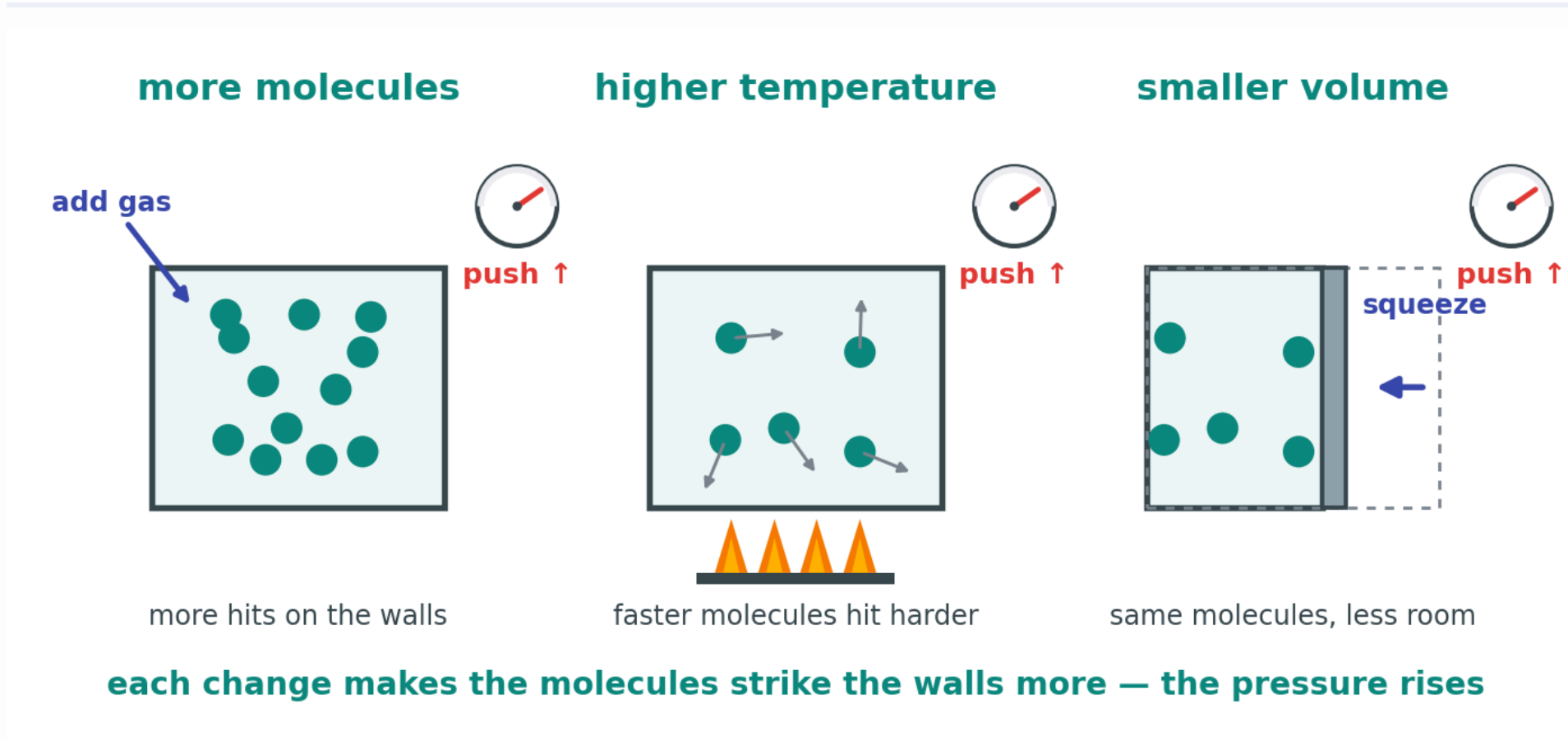
Pressure is force per unit area

Pressure measures how hard the gas pushes on each unit of wall area. The same total push, spread over more area, gives less pressure.

$$\text{pressure} = \text{force} \div \text{area}$$

more molecules, or faster ones, means more force, and more pressure

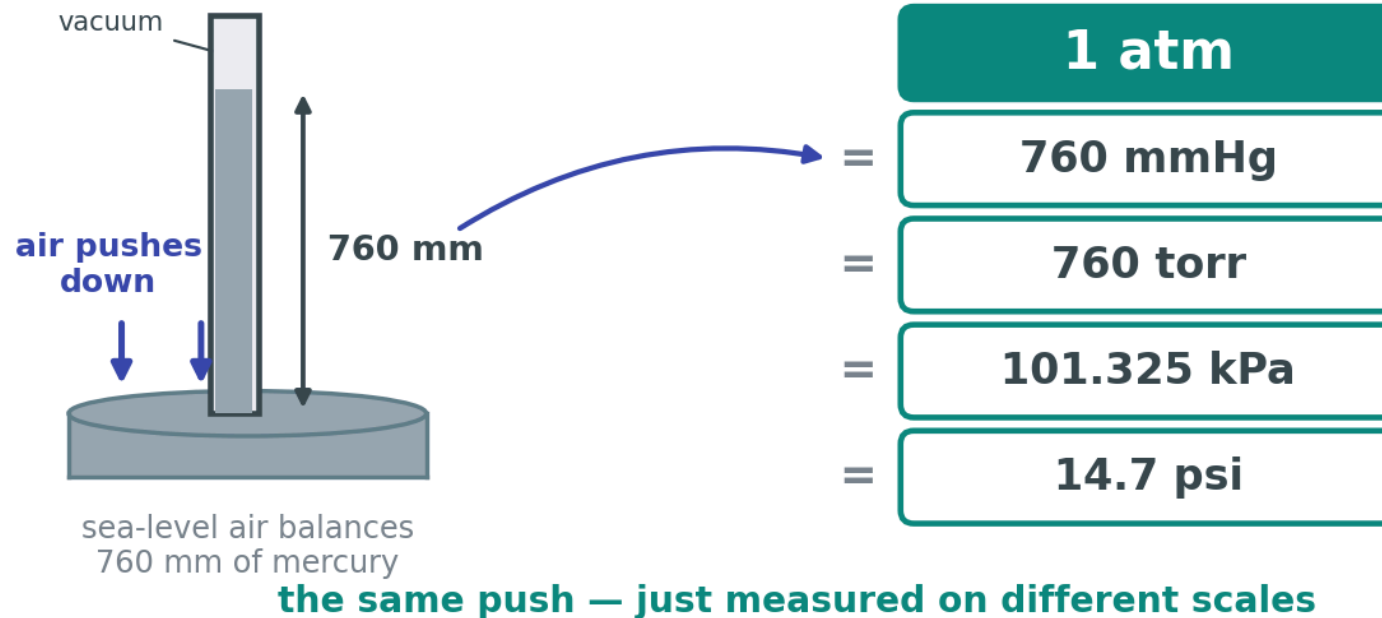
What raises the pressure



Three changes make the molecules strike the walls more: more molecules, a higher temperature, or a smaller volume. The exact amounts come from the gas laws.

The units of pressure

one pressure, five names



One pressure has many names. A barometer shows that sea-level air holds up 760 mm of mercury, and that pressure is called one atmosphere. Every unit below names the same push.



The method

1. **Write the given:** the pressure, its unit, and the wanted unit.
2. **Pick the factor that cancels the given unit:** it goes in the denominator.
3. **Multiply; repeat until the wanted unit survives.**
4. **Sense-check** the size and the surviving unit.



Worked example 1: cylinder gauge

Step 1 • Write the given

$$1 \text{ atm} = 760 \text{ mmHg}$$

given: 2.50 atm • wanted: mmHg

A compressed-gas cylinder gauge reads **2.50 atm**. Express the pressure in mmHg.

A common first attempt writes the factor as 1 atm over 760 mmHg. Test it.

✓ Worked example 1: solution

1 atm = 760 mmHg
given: 2.50 atm · wanted: mmHg

A common first attempt

Step 2 · Pick the factor that cancels its unit

Step 3 · Multiply; repeat until the wanted unit survives

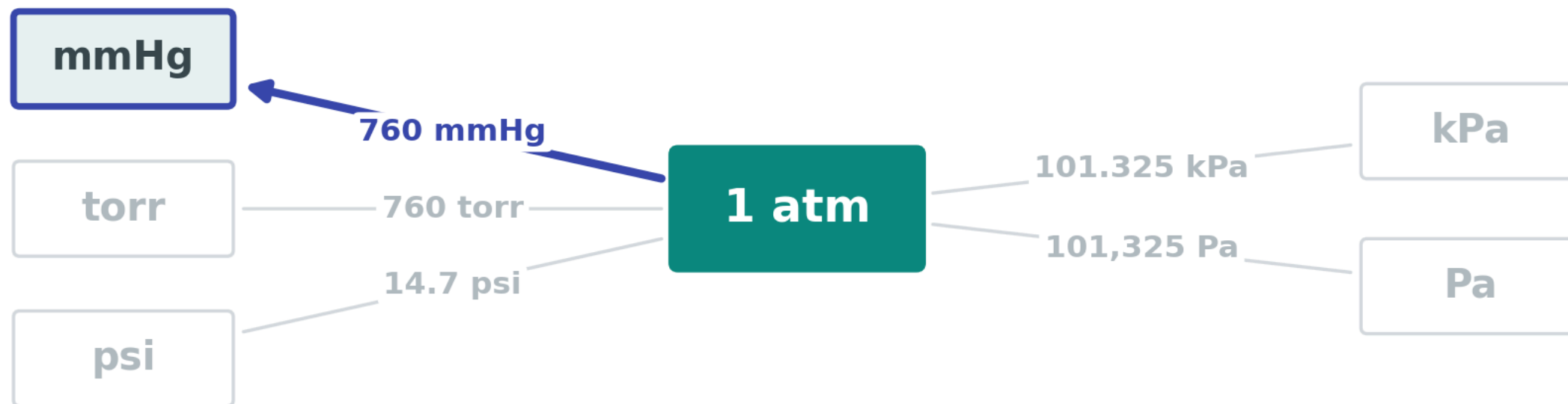
Step 4 · Sense-check

$$2.50 \text{ atm} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} = 1900 \text{ mmHg}$$

 **Sense check:** A millimeter of mercury is a far smaller unit than an atmosphere, so the count grows: 2.50 → 1900. The pressure itself is unchanged. ✓

✓ Worked example 1: the route on the map

1 atm = 760 mmHg
given: 2.50 atm · found: 1900 mmHg



🧠 **Sense check:** The given already sits at the center of the map, so one arrow out reaches mmHg: one conversion factor. ✓

Take-home: only the orientation that cancels

Do: the given unit in the denominator, so it cancels.

$$2.50 \text{ atm} \times (760 \text{ mmHg} / 1 \text{ atm}) = 1900 \text{ mmHg}$$

atm cancels · mmHg survives ✓

Do not: the given unit on top. Nothing cancels; the factor is upside down.

$$2.50 \text{ atm} \times (1 \text{ atm} / 760 \text{ mmHg}) = 3.29 \times 10^{-3} \text{ atm}^2/\text{mmHg}$$

no unit cancels · flip the factor ✗

Your turn: a storm barometer

$$1 \text{ atm} = 101.325 \text{ kPa}$$

given: 0.850 atm · wanted: kPa

A barometer reads **0.850 atm** as a storm moves in. Fill in the factor, then compute.

$$0.850 \text{ atm} \times \frac{\text{ kPa }}{\text{ atm }} = \text{ kPa }$$

Fill the factor from $1 \text{ atm} = 101.325 \text{ kPa}$, then compute.

$$0.850 \text{ atm} \times \frac{101.325 \text{ kPa }}{1 \text{ atm }} = \mathbf{86.1 \text{ kPa}}$$

Where this goes wrong

Grabbing the wrong equivalence. Each unit pairs with atm through its own number: 760 mmHg, 760 torr, 101.325 kPa, or 14.7 psi. Reading 44.1 psi with the mmHg number, $44.1 / 760 = 0.0580$ atm, is wrong; its own number gives $44.1 / 14.7 = 3.00$ atm. Match the unit to its own equivalence.

Stopping mid-plan. The plan $\text{kPa} \rightarrow \text{atm} \rightarrow \text{torr}$ has two arrows. Stopping after one leaves 202.6 kPa as 2.00 atm: atmospheres, not the wanted torr. The chain ends at 1520 torr.

Treating torr and mmHg as different. They are the same size: $1 \text{ torr} = 1 \text{ mmHg}$, and both count 760 to one atmosphere. A reading already in torr needs no conversion to reach mmHg.



Practice 1: a bicycle tire

$$1 \text{ atm} = 14.7 \text{ psi}$$

given: 88.2 psi · wanted: atm

A road-bike tire is pumped to **88.2 psi**. What is this pressure in atmospheres?

- A. 1.30×10^3
- B. 0.116
- C. 6.00
- D. 0.870

✓ Practice 1 answer: C

$$1 \text{ atm} = 14.7 \text{ psi}$$

given: 88.2 psi · wanted: atm

$$88.2 \text{ psi} \times \frac{1 \text{ atm}}{14.7 \text{ psi}} = 6.00 \text{ atm} \rightarrow \text{answer C}$$

A inverted the factor: $88.2 \times 14.7 = 1.30 \times 10^3$, and no unit cancels; the units come out psi^2/atm , not atm. B used mercury's 760 in place of psi's 14.7: $88.2 / 760 = 0.116$. D used the kPa number 101.325: $88.2 / 101.325 = 0.870$.

 **Sense check:** A tire runs well above room air, so more than 1 atm: 6.00. A count in psi reads larger than the same pressure in atm. ✓

Worked example 2: hospital regulator

Step 1 • Write the given

$$1 \text{ atm} = 101.325 \text{ kPa} \cdot 1 \text{ atm} = 760 \text{ torr}$$

given: 202.6 kPa • wanted: torr • plan: kPa → atm → torr

A hospital oxygen regulator reads **202.6 kPa**. Express the pressure in torr.

No single equality links kPa to torr. The plan runs through atmospheres: kPa → atm → torr.

✓ Worked example 2: solution

1 atm = 101.325 kPa · 1 atm = 760 torr
 given: 202.6 kPa · wanted: torr · plan: kPa → atm → torr

Step 2 · Pick the factor that cancels its unit

$$202.6 \text{ kPa} \times \frac{1 \text{ atm}}{101.325 \text{ kPa}} = 2.00 \text{ atm}$$

Step 3 · Multiply; repeat until the wanted unit survives

Step 4 · Sense-check

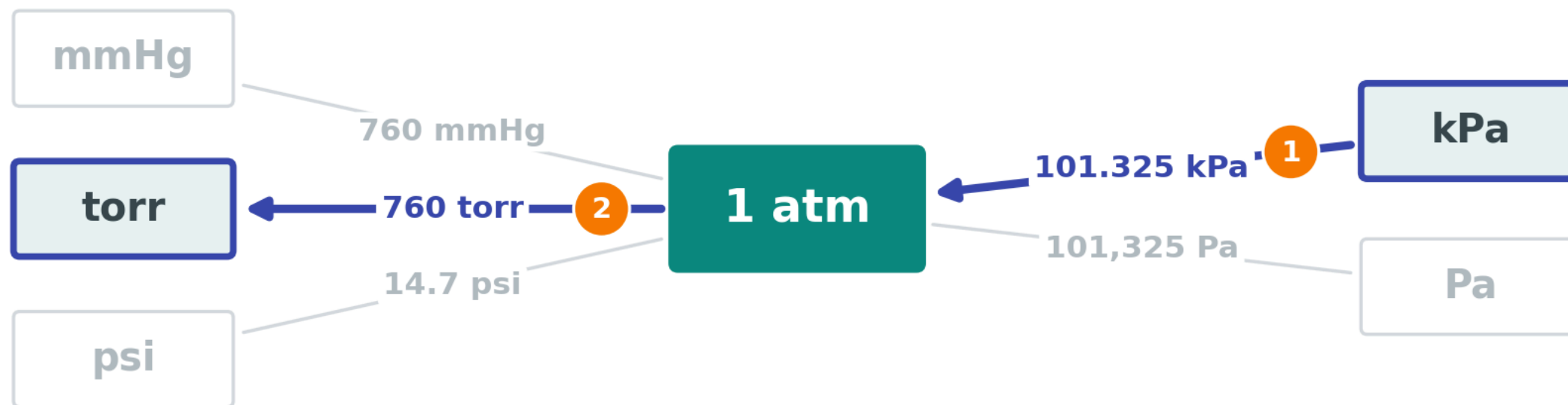
$$202.6 \text{ kPa} \times \frac{1 \text{ atm}}{101.325 \text{ kPa}} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 1520 \text{ torr}$$

🧠 **Sense check:** 202.6 kPa is close to 2 atm, and each atmosphere is 760 torr, so the answer lands near 1520 torr. ✓

✓ Worked example 2: the route on the map

$$1 \text{ atm} = 101.325 \text{ kPa} \cdot 1 \text{ atm} = 760 \text{ torr}$$

given: 202.6 kPa · found: 1520 torr



🧠 **Sense check:** No spoke joins kPa to torr. Arrow 1 goes into atm and arrow 2 comes out: two conversion factors. ✓

Practice 2: a pressure cooker

$$1 \text{ atm} = 760 \text{ torr} \cdot 1 \text{ atm} = 101.325 \text{ kPa}$$

given: 758 torr at the start · 162.0 kPa when hot · wanted: the rise in atm

A sealed pressure cooker starts at **758 torr**. On the stove it reaches **162.0 kPa**. By how many atmospheres did the pressure rise?

- A. 1.60
- B. 0.601
- C. 60.9
- D. -0.601

✓ Practice 2 answer: B

$$1 \text{ atm} = 760 \text{ torr} \cdot 1 \text{ atm} = 101.325 \text{ kPa}$$

given: 758 torr → 162.0 kPa · two readings, two units: both go to atm before they subtract

$$758 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.9974 \text{ atm} \cdot 162.0 \text{ kPa} \times \frac{1 \text{ atm}}{101.325 \text{ kPa}} = 1.5988 \text{ atm}$$

$$\text{rise} = 1.5988 \text{ atm} - 0.9974 \text{ atm} = \mathbf{0.601 \text{ atm} \rightarrow \text{answer B}}$$

A stopped before subtracting: 1.60 atm is the final reading, not the rise. C subtracted in kPa and skipped the last hop: $162.0 - 101.06 = 60.9$, still kPa. D subtracted backward: $0.9974 - 1.5988 = -0.601$, a fall, though the pressure rose.

 **Sense check:** From about 1 atm to about 1.6 atm: a rise of a little over half an atmosphere. ✓

Check yourself

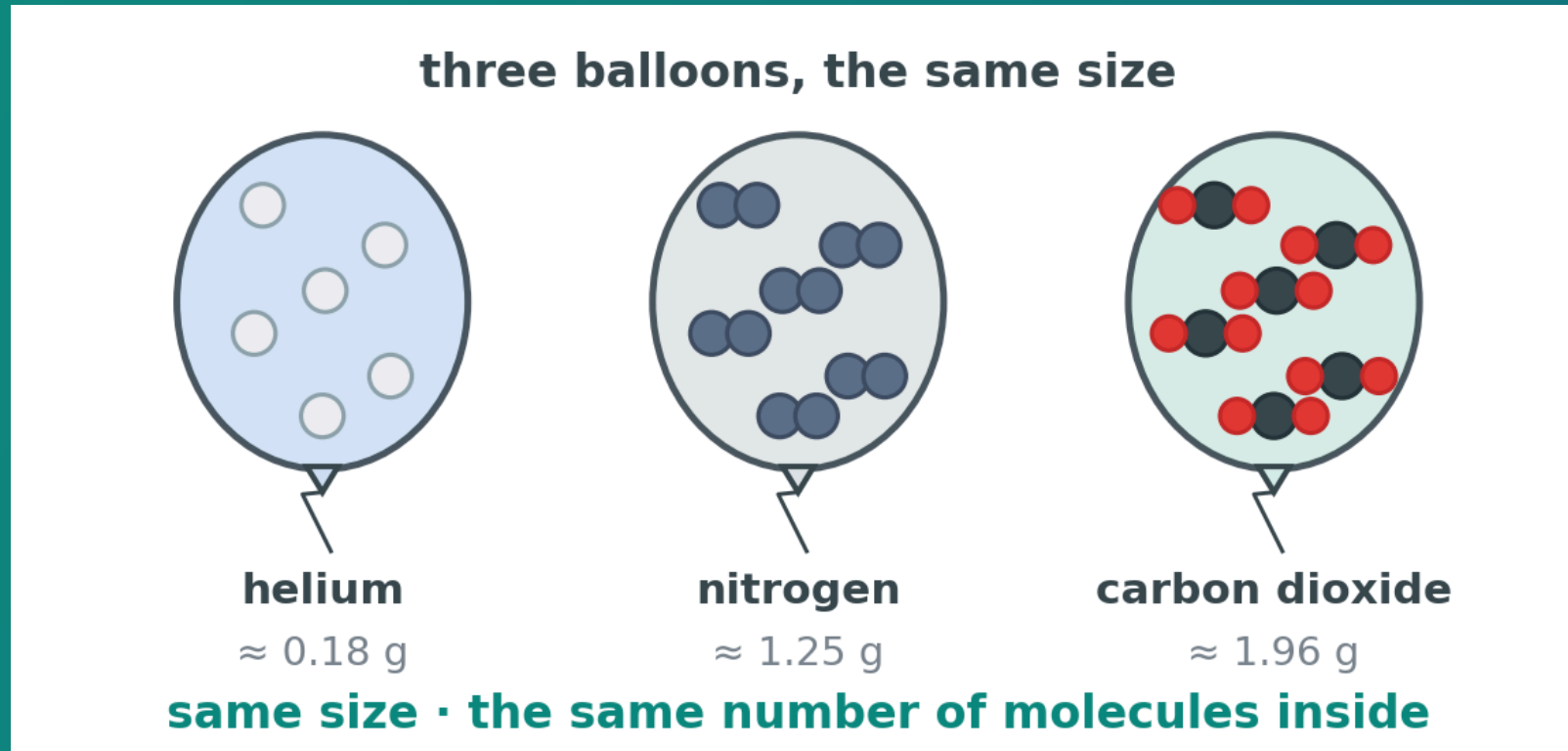
1. From **1 atm = 760 mmHg**, write both conversion factors. Which one converts a reading of 950 mmHg to atm?
2. A sealed rigid can is thrown on a fire. Do its molecules strike the walls harder or softer, and does the pressure rise or fall?

Pressure, volume, temperature, and amount all move together. The gas laws turn those relationships into equations: the combined gas law, then the ideal gas law. Every one reads pressure in one of these units.

2 • Molar Volume at STP

Use the molar volume of a gas at STP, 22.4 L/mol, to convert between moles and liters, and chain it with molar mass to go from grams of a gas to its volume.

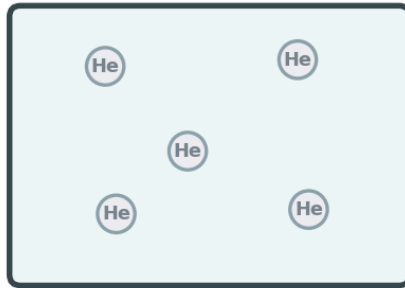
What a balloon holds



Fill three identical balloons: one with helium, one with nitrogen, one with carbon dioxide. The gases weigh very different amounts. Yet each balloon holds the same number of molecules.

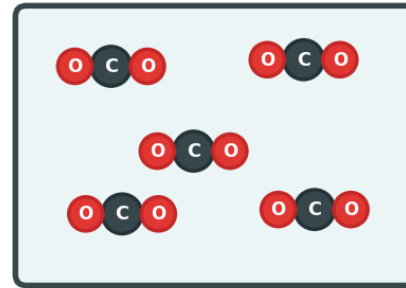
Equal volumes, equal counts

same temperature and pressure · same volume



helium

1 mol = 4.00 g



carbon dioxide

1 mol = 44.01 g

same number of molecules → same volume, whichever gas it is

Gas molecules are specks in mostly empty space, so identity barely matters. At the same temperature and pressure, equal volumes hold equal numbers of molecules. One mole of any gas fills the same volume.

Avogadro's law: $V \propto n \rightarrow V_1/n_1 = V_2/n_2$

fixed T and P · double the moles, double the volume

STP fixes the conditions

1 mol of any gas = 22.4 L at STP

STP: 0 °C = 273.15 K and 1 atm · molar volume = 22.4 L/mol

"Same conditions" needs a reference. Chemists use STP: 0 °C and 1 atm. At STP, one mole of any gas occupies 22.4 liters. That shared volume is the molar volume.

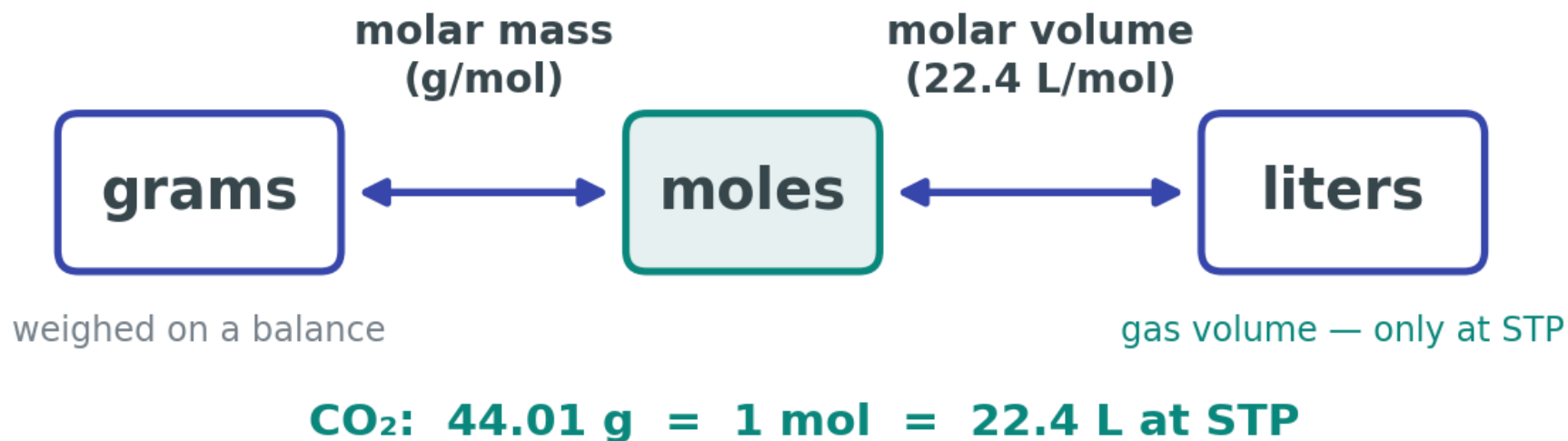
The molar volume is an equality

The molar volume, $22.4 \text{ L} = 1 \text{ mol}$, is an equality. Every equality gives two conversion factors. This one converts between moles and volume:

$$\frac{22.4 \text{ L}}{1 \text{ mol}} \text{ or } \frac{1 \text{ mol}}{22.4 \text{ L}}$$

Write it so the given unit cancels.

One route: grams, moles, liters



Molar mass links grams to moles. Molar volume links moles to liters at STP. A gas mass becomes a gas volume by passing through moles.



The method

1. **Confirm STP.** 22.4 L/mol applies only at 0 °C, 1 atm.
2. **Map the route:** grams → moles → liters.
3. **Orient each factor** so the unit to cancel sits underneath.
4. **Multiply and check** that only the wanted unit survives.



Worked example 1: moles to liters

Step 1 • Confirm STP

$$22.4 \text{ L} = 1 \text{ mol (at STP)}$$

given: 0.750 mol He • wanted: volume in L

A weather balloon is filled with **0.750 mol of helium** at STP. What volume does the gas occupy?

Set it up so the given moles cancel.

✓ Worked example 1: solution

22.4 L = 1 mol (at STP)

given: 0.750 mol He · wanted: volume in L

Step 2 · Map the route

Step 3 · Orient each factor

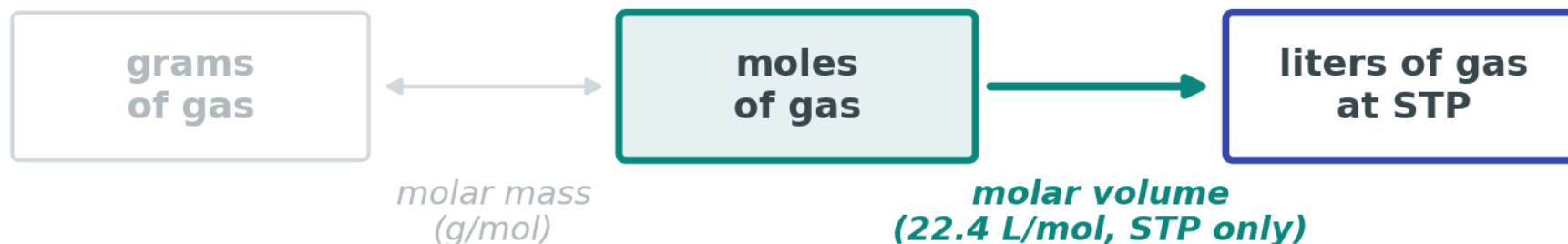
Step 4 · Multiply and check

$$0.750 \cancel{\text{ mol He}} \times \frac{22.4 \text{ L}}{1 \cancel{\text{ mol He}}} = 16.8 \text{ L}$$

🧠 **Sense check:** One mole of any gas is 22.4 L at STP. Three-quarters of a mole is three-quarters of that: 16.8 L. ✓

✓ Worked example 1: the route on the map

22.4 L = 1 mol (at STP)
given: 0.750 mol He · found: 16.8 L



moles → liters: one factor

🧠 **Sense check:** Moles in, liters out: one arrow, the molar volume. The grams box stays grey, because no mass entered the problem. ✓



Worked example 2: liters to moles

Step 1 • Confirm STP

$$22.4 \text{ L} = 1 \text{ mol (at STP)}$$

given: 5.60 L N₂ • wanted: mol N₂

A rigid **5.60 L** cylinder holds nitrogen at STP. How many moles of N₂ is that?

A common first attempt multiplies by the molar volume as written, 22.4 L over 1 mol. Test the units.

✓ Worked example 2: solution

22.4 L = 1 mol (at STP)
given: 5.60 L N₂ · wanted: mol N₂

A common first attempt

$$5.60 \text{ L N}_2 \times \frac{22.4 \text{ L}}{1 \text{ mol N}_2} = 125 \text{ L}^2/\text{mol} \times$$

Step 2 · Map the route

Step 3 · Orient each factor

Step 4 · Multiply and check

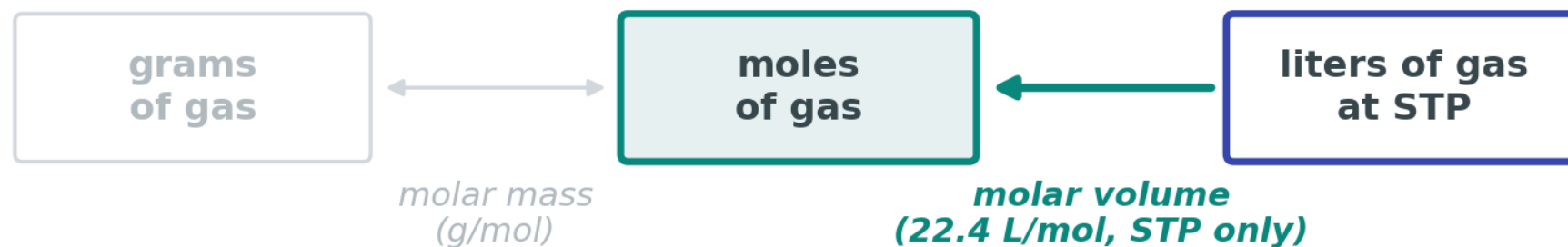
$$5.60 \text{ L} \times \frac{1 \text{ mol N}_2}{22.4 \text{ L}} = 0.250 \text{ mol N}_2$$

 **Sense check:** One mole fills 22.4 L. The cylinder holds a quarter of that volume, so it holds a quarter of a mole: 0.250. ✓

✓ Worked example 2: the route on the map

$$22.4 \text{ L} = 1 \text{ mol (at STP)}$$

given: 5.60 L N₂ · found: 0.250 mol N₂



liters → moles: one factor

🧠 **Sense check:** The same arrow, run the other way: liters in, moles out. The factor flips, so liters sit underneath. ✓

Take-home: which way the factor goes

For liters, put moles underneath so moles cancel. For moles, put liters underneath so liters cancel.

$$5.60 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = \mathbf{0.250 \text{ mol} \checkmark}$$

The upside-down factor cancels nothing and leaves units no gas volume carries.

$$5.60 \text{ L} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 125 \text{ L}^2/\text{mol} \times$$

Your turn: argon

22.4 L = 1 mol (at STP)
given: 0.400 mol Ar · wanted: L

A **0.400 mol** sample of argon sits in a flask at STP.

$$0.400 \text{ mol Ar} \times \frac{\text{ } \text{L}}{\text{ } \text{mol Ar}} = \text{ } \text{L}$$

Fill the factor so the given moles cancel, then compute.

$$0.400 \text{ mol Ar} \times \frac{22.4 \text{ L}}{1 \text{ mol Ar}} = \mathbf{8.96 \text{ L}}$$

Where this goes wrong

22.4 L = 1 mol (at STP)
given: 5.60 L N₂ · wanted: mol N₂

The molar volume upside down. $5.60 \text{ L} \times (22.4 \text{ L} / 1 \text{ mol}) = 125 \text{ L}^2/\text{mol}$. No unit cancels, and no gas volume carries L²/mol. Put liters underneath: $5.60 \text{ L} \times (1 \text{ mol} / 22.4 \text{ L}) = 0.250 \text{ mol}$.

Using 22.4 L/mol away from STP. The molar volume is 22.4 L/mol only at 0 °C and 1 atm. At warmer or higher-pressure conditions the same moles fill a different volume, and 22.4 no longer applies.

Grabbing molar mass instead. Molar mass, in g/mol, converts grams and moles. Molar volume, 22.4 L/mol, converts moles and liters. Reaching for the wrong one cancels the wrong unit.



Practice 1

22.4 L = 1 mol (at STP)

given: 3.36 L O₂ · wanted: mol O₂

An anesthesia line delivers **3.36 L of oxygen** at STP. How many moles of O₂ is that?

- A. 0.105
- B. 108
- C. 75.3
- D. 0.150

✓ Practice 1 answer: D

22.4 L = 1 mol (at STP)

given: 3.36 L O₂ · wanted: mol O₂

$$3.36 \cancel{\text{L O}_2} \times \frac{1 \text{ mol O}_2}{22.4 \cancel{\text{L O}_2}} = 0.150 \text{ mol O}_2 \rightarrow \text{answer D}$$

A reached for the molar mass, 32.00 g/mol, instead of the molar volume: $3.36 / 32.00 = 0.105$, and grams never entered the problem. C left the molar volume upside down: $3.36 \times 22.4 = 75.3$, with units L²/mol. B used the molar mass upside down: $3.36 \times 32.00 = 108$. Only 22.4 L/mol, with liters underneath, cancels liters.

🧠 **Sense check:** One mole fills 22.4 L, and 3.36 L is well under a mole, so the answer is a small fraction: 0.150 mol. ✓



Worked example 3: grams to liters

Step 1 • Confirm STP

$$\text{CO}_2: 12.01 + 2(16.00) = 44.01 \text{ g/mol}$$

given: 22.0 g CO₂ at STP • wanted: L • molar volume 22.4 L/mol

A dry-ice pellet sublimates into **22.0 g of CO₂ gas** at STP. What volume does it fill?

No single equality links grams to liters. Build the chain so each unit cancels the one before.

✓ Worked example 3: solution

CO₂: 44.01 g/mol · 22.4 L = 1 mol at STP
given: 22.0 g CO₂ · wanted: L

Step 2 · Map the route

Step 3 · Orient each factor

Step 4 · Multiply and check

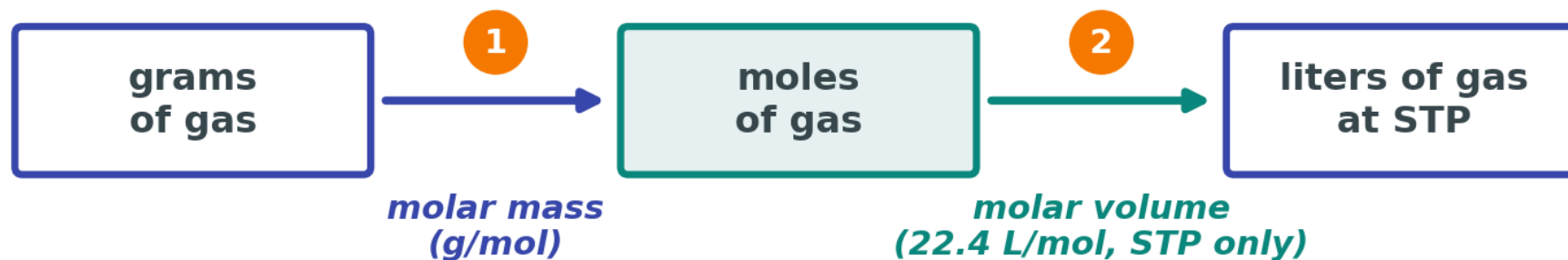
$$22.0 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{22.4 \text{ L}}{1 \text{ mol CO}_2} = 11.2 \text{ L}$$

🧠 **Sense check:** The pellet is 0.500 mol of CO₂. Half a mole fills half of 22.4 L: 11.2 L. ✓

✓ Worked example 3: the route on the map

CO_2 : $44.01 \text{ g/mol} \cdot 22.4 \text{ L} = 1 \text{ mol at STP}$

given: 22.0 g CO_2 · found: 11.2 L



grams → moles → liters: two factors

🧠 **Sense check:** Two arrows: the molar mass reaches moles, then the molar volume reaches liters. ✓



Practice 2

22.4 L = 1 mol at STP

given: 6.72 L SO₂ at STP · wanted: g SO₂

A smokestack test captures **6.72 L of sulfur dioxide** at STP. What mass of SO₂, in grams, is that?

- A. 19.2
- B. 0.300
- C. 4.68×10^{-3}
- D. 2.35

✓ Practice 2 answer: A

SO_2 : $32.07 + 2(16.00) = 64.07 \text{ g/mol} \cdot 22.4 \text{ L} = 1 \text{ mol at STP}$

given: 6.72 L SO_2 · wanted: g · route: $\text{L} \rightarrow \text{mol} \rightarrow \text{g}$

$$6.72 \text{ L SO}_2 \times \frac{1 \text{ mol SO}_2}{22.4 \text{ L SO}_2} \times \frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} = 19.2 \text{ g SO}_2 \rightarrow \text{answer A}$$

B stopped at moles: $6.72 / 22.4 = 0.300 \text{ mol}$, an amount, not a mass. C flipped the molar mass: $0.300 / 64.07 = 4.68 \times 10^{-3}$, grams underneath. D ran the grams-to-liters chain on a volume, both factors upside down: $6.72 / 64.07 \times 22.4 = 2.35$.

🧠 **Sense check:** 0.300 mol of a gas at 64 g per mole is about a third of 64 : 19.2 g . ✓

Check yourself

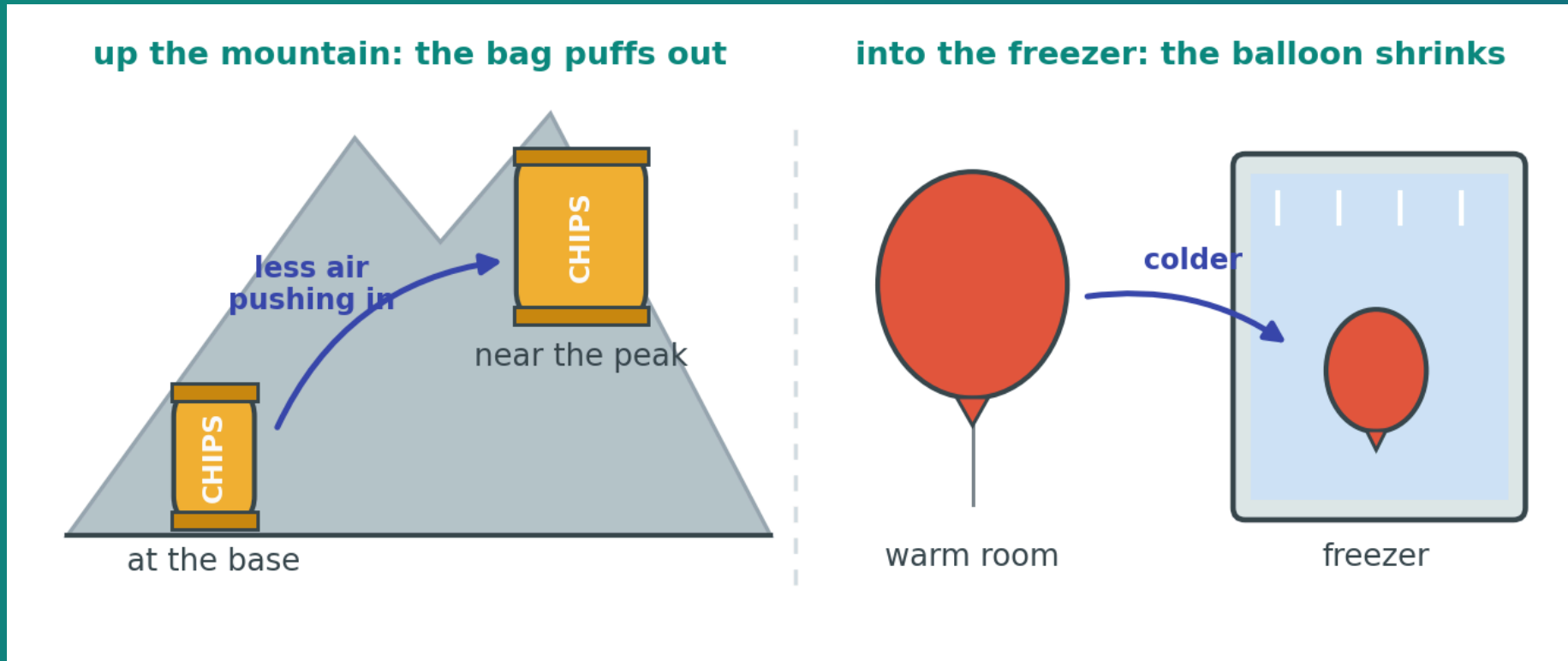
1. From **22.4 L = 1 mol at STP**, write both conversion factors. Which one converts 0.300 mol of a gas to liters, and which converts 44.8 L of a gas to moles?
2. One flask holds 1.00 mol of helium; another holds 1.00 mol of carbon dioxide. Both are at STP. Which flask has the greater volume? Which holds the greater mass?

Molar volume is the STP shortcut. Away from STP, a gas volume follows the ideal gas law, $PV = nRT$, where n is the mole count you find here.

3 · The Combined Gas Law

Use $P_1V_1/T_1 = P_2V_2/T_2$ to find any one final pressure, volume, or temperature of a fixed amount of gas, converting every temperature to kelvin before the ratios go in.

A bag that puffs, a balloon that shrinks



A sealed chip bag swells on a mountain road. A balloon in the freezer sags. Squeeze, heat, or cool a trapped gas, and its volume changes.

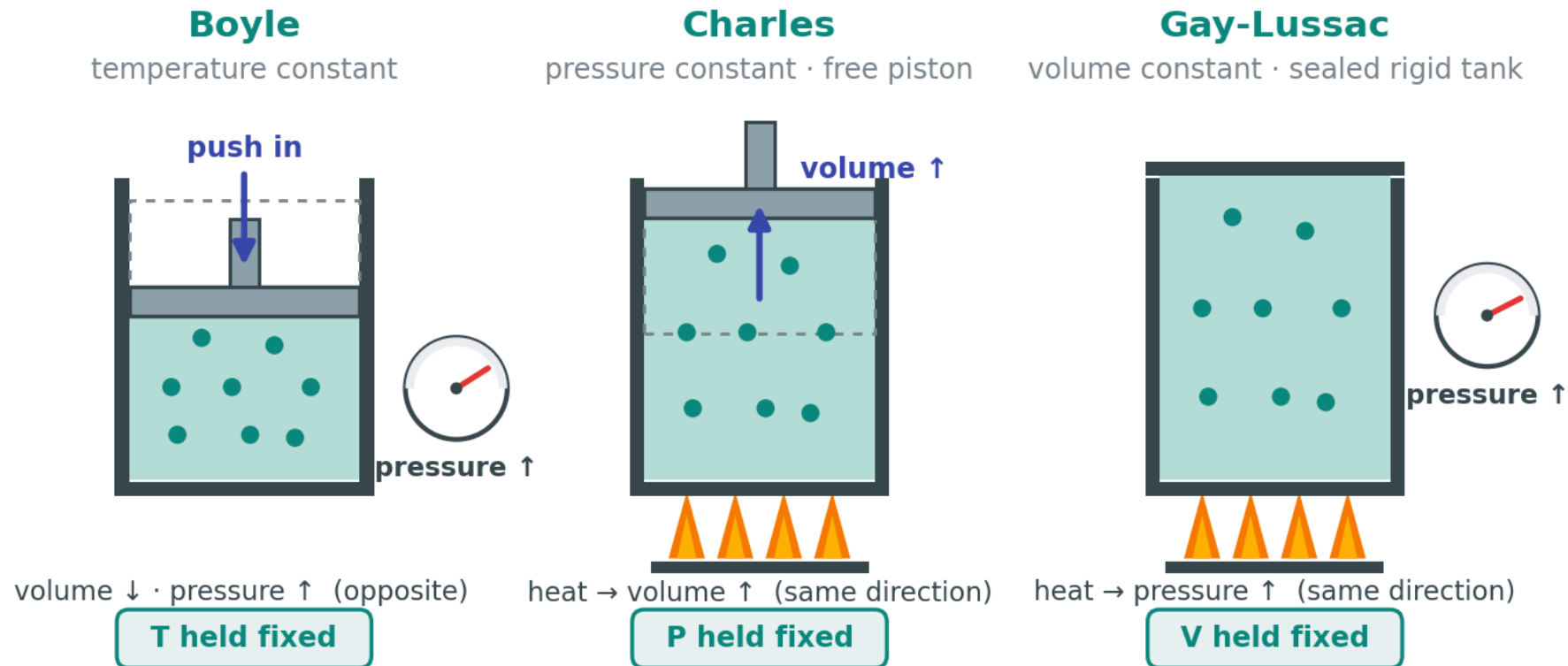
A fixed gas keeps PV/T constant

$$PV / T = \text{constant} \rightarrow P_1V_1 / T_1 = P_2V_2 / T_2$$

a fixed amount of gas · P in atm · V in L · T in kelvin

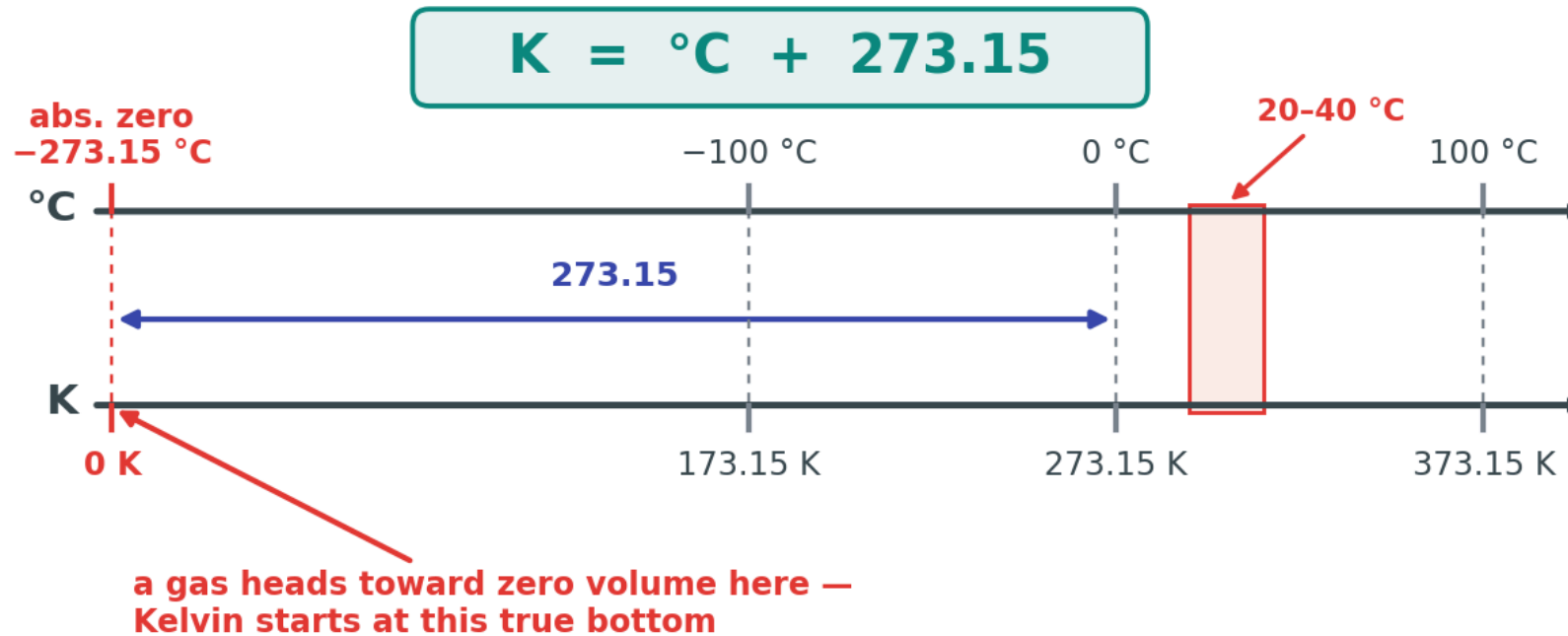
For a sealed sample of gas, PV/T does not change. Squeeze it, heat it, or cool it: pressure, volume, and temperature shift together to hold PV/T fixed. Change two, the third follows.

Hold one variable fixed



Freeze one variable and the law simplifies. At constant temperature, pressure and volume move in opposite directions (Boyle). At constant pressure, volume rises with temperature (Charles). At constant volume, pressure rises with temperature (Gay-Lussac).

Temperature must be in kelvin



20 °C and 40 °C are not double: as 293.15 K and 313.15 K they sit about 7% apart. A gas responds to the Kelvin ratio.

Gas-law ratios only work on a scale that starts at absolute zero. Kelvin does; Celsius does not. Convert first: $K = ^\circ\text{C} + 273.15$. A Celsius ratio can even return a negative volume.



The method

1. **Identify given and wanted.** Mark the unknown.
2. **Convert to kelvin.** Every temperature: $K = ^\circ C + 273.15$.
3. **Rearrange for the unknown.**
4. **Substitute and check.** Did the volume move the sensible way?

Worked example 1: compressing a gas at constant temperature

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: $V_1 = 6.0 \text{ L}$ · $P_1 = 1.0 \text{ atm}$ · $P_2 = 3.0 \text{ atm}$ · temperature constant · wanted: V_2

A sealed cylinder holds **6.0 L of gas at 1.0 atm**. Held at constant temperature, it is compressed until the pressure reads **3.0 atm**. Find the new volume.

Identify the given and the wanted, and note what is held fixed.

✓ Worked example 1: solution

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: $V_1 = 6.0 \text{ L}$ · $P_1 = 1.0 \text{ atm}$ · $P_2 = 3.0 \text{ atm}$ · temperature constant · wanted: V_2

Step 1 · Identify given and wanted

Step 2 · Convert to kelvin

Step 3 · Rearrange for the unknown

$$P_1V_1 = P_2V_2 \rightarrow V_2 = V_1 \times \frac{P_1}{P_2}$$

Step 4 · Substitute and check

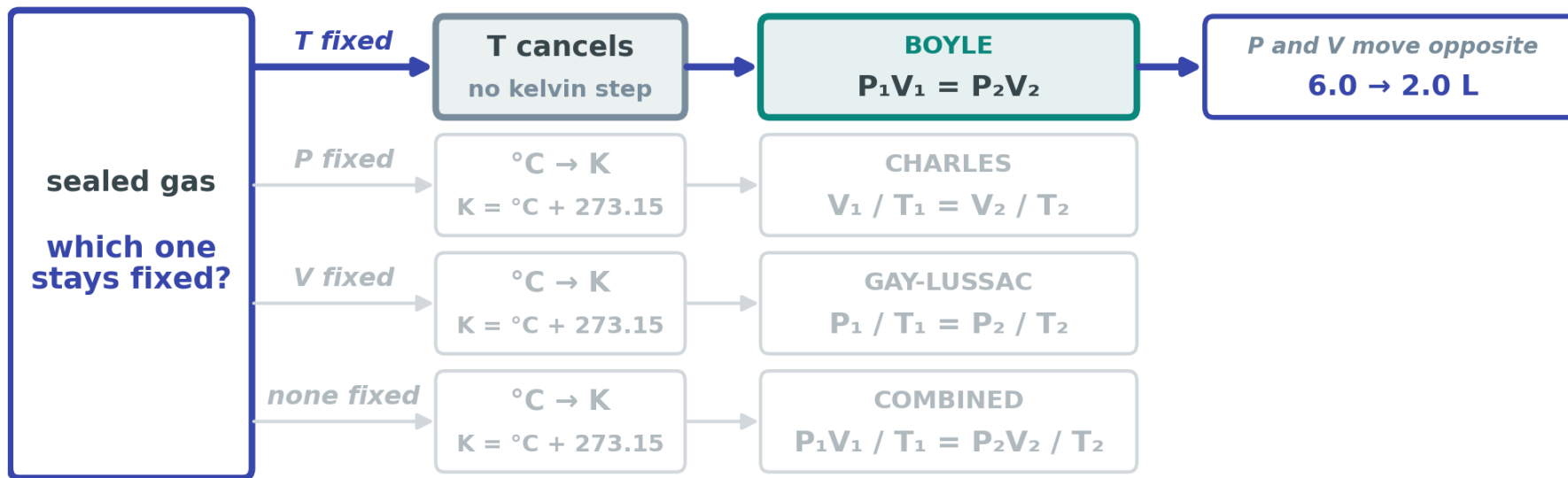
$$V_2 = 6.0 \text{ L} \times \frac{1.0 \text{ atm}}{3.0 \text{ atm}} = 2.0 \text{ L}$$

 **Sense check:** Pressure tripled at constant temperature, so the volume falls to a third: $6.0 \text{ L} \rightarrow 2.0 \text{ L}$. ✓

✓ Worked example 1: the route on the chart

$$V_2 = 6.0 \text{ L} \times 1.0 \text{ atm} / 3.0 \text{ atm} = 2.0 \text{ L}$$

T fixed · no kelvin step · found: $V_2 = 2.0 \text{ L}$



Sense check: Temperature held constant, so T cancels and the kelvin step drops out. Boyle's law alone does the work. ✓



Worked example 2: heating a gas at constant pressure

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: $V_1 = 3.00 \text{ L}$ · $25.0 \text{ }^\circ\text{C} \rightarrow 75.0 \text{ }^\circ\text{C}$ · pressure constant · wanted: V_2

A **3.00 L** gas sample warms from **25.0 °C to 75.0 °C** under a free-riding piston, so the pressure stays constant. Find the new volume.

List the given and the wanted, and mark what is held fixed.

✓ Worked example 2: solution

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: $V_1 = 3.00 \text{ L} \cdot 25.0 \text{ }^\circ\text{C} \rightarrow 75.0 \text{ }^\circ\text{C} \cdot \text{pressure constant} \cdot \text{wanted: } V_2$

Step 1 · Identify given and wanted

Step 2 · Convert to kelvin

Step 3 · Rearrange for the unknown

$$P \text{ constant} \rightarrow \frac{V_1}{T_1} = \frac{V_2}{T_2} \rightarrow V_2 = V_1 \times \frac{T_2}{T_1}$$

Step 4 · Substitute and check

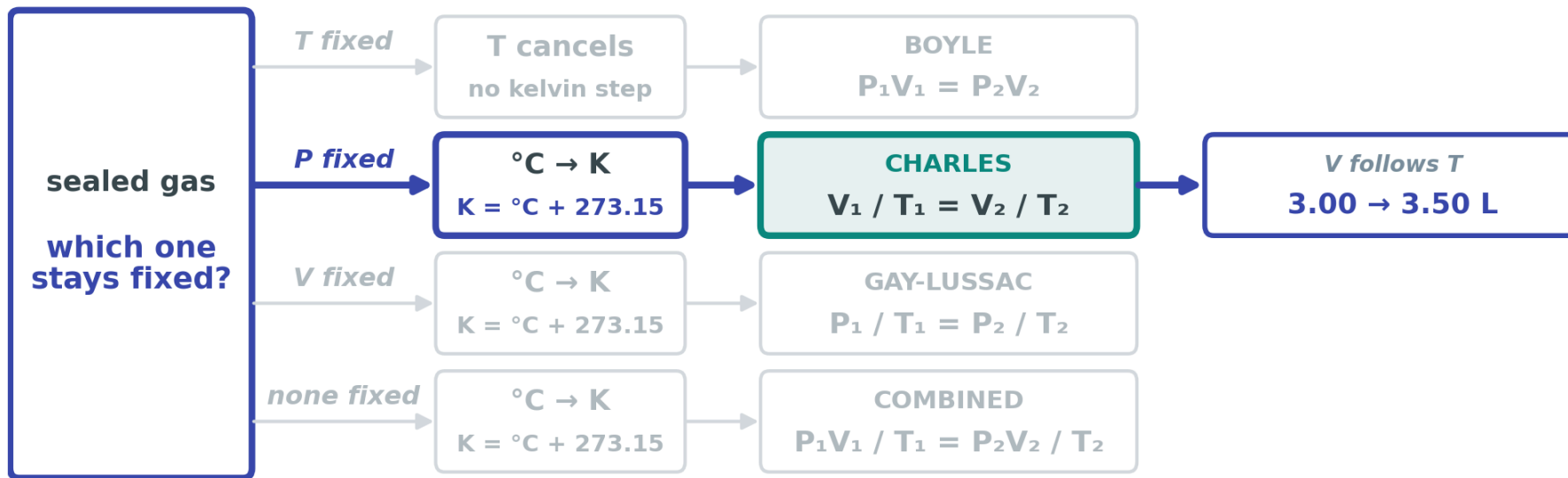
$$V_2 = 3.00 \text{ L} \times \frac{348.15 \text{ K}}{298.15 \text{ K}} = 3.50 \text{ L}$$

 **Sense check:** Warming expands a gas at constant pressure, so the volume grows: $3.00 \text{ L} \rightarrow 3.50 \text{ L}$. A raw Celsius ratio $75.0/25.0$ would wrongly triple it to 9.00 L . ✓

✓ Worked example 2: the route on the chart

$$V_2 = 3.00 \text{ L} \times 348.15 \text{ K} / 298.15 \text{ K} = 3.50 \text{ L}$$

P fixed · kelvin step used · found: $V_2 = 3.50 \text{ L}$



 **Sense check:** A free piston holds the pressure fixed. Temperature changed, so the route needs the kelvin step before Charles's law. ✓

Your turn: a sealed can heated at constant volume

$$P_1 / T_1 = P_2 / T_2$$

given: $P_1 = 2.00 \text{ atm}$ · $20.0 \text{ }^\circ\text{C} \rightarrow 120.0 \text{ }^\circ\text{C}$ · volume constant · wanted: P_2

$$P_2 = 2.00 \text{ atm} \times \frac{\text{ } \text{K}}{\text{ } \text{K}} = \text{ } \text{atm}$$

$$P_2 = 2.00 \text{ atm} \times \frac{393.15 \text{ K}}{293.15 \text{ K}} = \mathbf{2.68 \text{ atm}}$$

 **Sense check:** Volume is fixed, so heating drives the pressure up: $2.00 \text{ atm} \rightarrow 2.68 \text{ atm}$. This is why sealed cans warn against incineration. ✓

⚠ Where this goes wrong

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

$$4.00 \text{ L} \cdot 2.00 \text{ atm}, 27^\circ\text{C} \rightarrow 1.00 \text{ atm}, 127^\circ\text{C} \cdot \text{correct } V_2 = 10.7 \text{ L}$$

Leaving temperature in Celsius. Using 127/27 instead of 400.15/300.15 gives $4.00 \times (2.00/1.00) \times (127/27) = 37.6 \text{ L}$. A Celsius ratio exaggerates the change wildly. Convert first: $K = ^\circ\text{C} + 273.15$.

Inverting the pressure ratio. Pressure fell from 2.00 to 1.00 atm, so the gas expands. Writing $(1.00/2.00)$ gives 2.67 L, a shrinking gas. The old pressure goes on top: (P_1/P_2) .

Inverting the temperature ratio. Heating from 300.15 to 400.15 K expands the gas. Writing $(300.15/400.15)$ gives 6.00 L. The new temperature goes on top: (T_2/T_1) .

Holding a variable that actually changed. Treating temperature as constant uses pressure alone: $4.00 \times (2.00/1.00) = 8.00 \text{ L}$. Both P and T changed here, so both ratios belong in the setup.

 Practice 1

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: $V_1 = 2.50 \text{ L} \cdot 17^\circ\text{C} \rightarrow 77^\circ\text{C}$ · pressure constant · wanted: V_2

A **2.50 L** balloon at **17 °C** is carried into a warm greenhouse and heated to **77 °C** at constant pressure. What is its new volume, in L?

- A. 0.552
- B. 2.07
- C. 3.02
- D. 11.3

✓ Practice 1 answer: C

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

$$\text{pressure constant} \cdot T_1 = 17 + 273.15 = 290.15 \text{ K} \cdot T_2 = 77 + 273.15 = 350.15 \text{ K}$$

$$V_2 = 2.50 \text{ L} \times \frac{350.15 \text{ K}}{290.15 \text{ K}} = 3.02 \text{ L} \rightarrow \text{answer C}$$

A left the temperatures in Celsius and inverted them: $2.50 \times (17/77) = 0.552 \text{ L}$. B inverted the kelvin ratio: $2.50 \times (290.15/350.15) = 2.07 \text{ L}$, a shrinking gas though it was heated. D left the temperatures in Celsius: $2.50 \times (77/17) = 11.3 \text{ L}$, a ratio that far overstates a 60-degree warming.

 **Sense check:** Heating at constant pressure expands the gas. Kelvin rose by about a fifth ($350.15 \div 290.15 = 1.21$), so the volume does too: $2.50 \text{ L} \rightarrow 3.02 \text{ L}$. ✓

Worked example 3: a balloon rising through the atmosphere

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: $V_1 = 2.00 \text{ L} \cdot 1.00 \text{ atm}, 27 \text{ }^\circ\text{C} \rightarrow 0.400 \text{ atm}, -23 \text{ }^\circ\text{C} \cdot$ wanted: V_2

A **2.00 L** helium balloon leaves the ground at **1.00 atm and 27 °C**. High up, the air is thinner and colder: **0.400 atm and -23 °C**. What volume does the gas reach?

A common first attempt: put the temperatures straight in as Celsius. Test the result.

✓ Worked example 3: solution

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: $V_1 = 2.00 \text{ L} \cdot 1.00 \text{ atm}, 27 \text{ }^\circ\text{C} \rightarrow 0.400 \text{ atm}, -23 \text{ }^\circ\text{C}$ · wanted: V_2

A common first attempt

$$V_2 = 2.00 \text{ L} \times \frac{1.00 \text{ atm}}{0.400 \text{ atm}} \times \frac{-23 \text{ }^\circ\text{C}}{27 \text{ }^\circ\text{C}} = -4.26 \text{ L} \times$$

Step 1 · Identify given and wanted

Step 2 · Convert to kelvin

$$T_1 = 27 + 273.15 = 300.15 \text{ K}. T_2 = -23 + 273.15 = 250.15 \text{ K}.$$

 **Sense check:** Above absolute zero every kelvin temperature is positive, so the ratio $250.15/300.15$ stays positive and just under 1: the volume will be real, and cooling trims it only a little. ✓

✓ Worked example 3: the volume at altitude

$$P_1 V_1 / T_1 = P_2 V_2 / T_2$$

$$V_1 = 2.00 \text{ L} \cdot P_1 = 1.00 \text{ atm}, P_2 = 0.400 \text{ atm} \cdot T_1 = 300.15 \text{ K}, T_2 = 250.15 \text{ K}$$

Step 3 • Rearrange for the unknown

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

Step 4 • Substitute and check

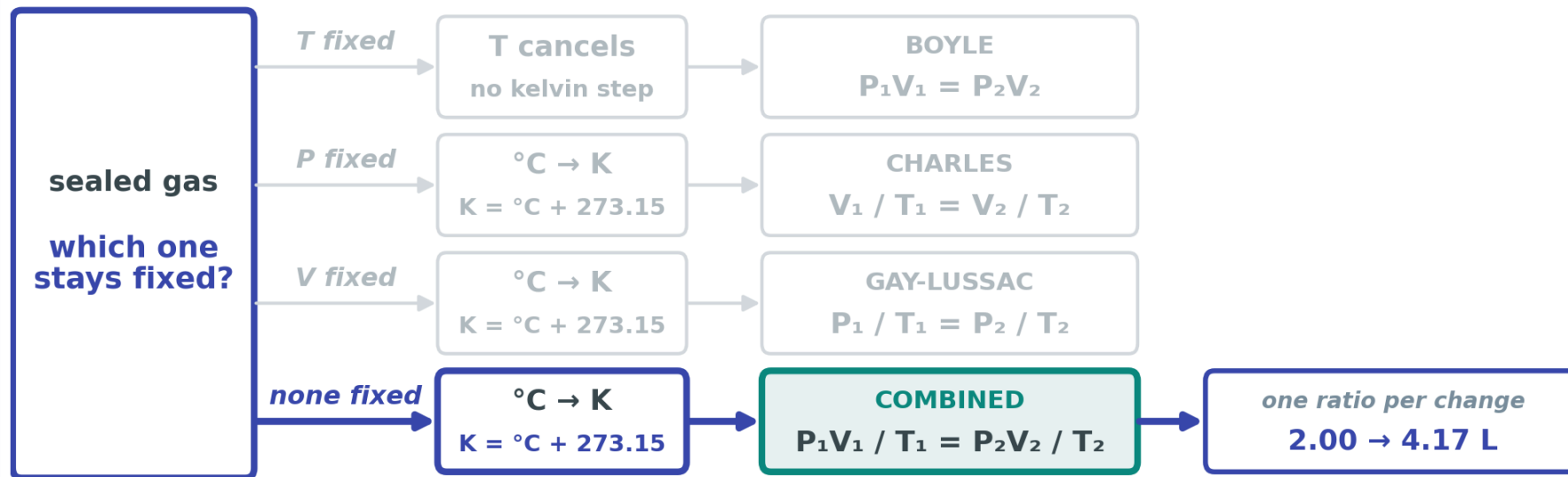
$$V_2 = 2.00 \text{ L} \times \frac{1.00 \text{ atm}}{0.400 \text{ atm}} \times \frac{250.15 \text{ K}}{300.15 \text{ K}} = 4.17 \text{ L}$$

 **Sense check:** Pressure dropped to less than half, expanding the gas, while the cooling pulled back only a little. Net result: 2.00 L → 4.17 L, a balloon that swells as it climbs. ✓

✓ Worked example 3: the route on the chart

$$V_2 = 2.00 \text{ L} \times (1.00 \text{ atm} / 0.400 \text{ atm}) \times (250.15 \text{ K} / 300.15 \text{ K}) = 4.17 \text{ L}$$

nothing fixed · kelvin step used · found: $V_2 = 4.17 \text{ L}$



 **Sense check:** Pressure and temperature both changed, so neither cancels. The combined law carries one ratio for each change. ✓

Take-home: kelvin, always

in Celsius: $-23 / 27 \rightarrow V_2 = -4.26 \text{ L}$

a negative volume: impossible ✕

in kelvin: $250.15 / 300.15 \rightarrow V_2 = 4.17 \text{ L}$

a smaller, positive volume ✓

A gas-law ratio taken in Celsius is meaningless, and below 0°C it turns negative: a volume cannot. Convert every temperature to kelvin before any ratio: $\text{K} = ^\circ\text{C} + 273.15$.

 Practice 2

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: 16.0 L · 0.980 atm · 21.0 °C → 40.0 L · 0.300 atm · wanted: T_2 in °C

At launch, a weather balloon holds **16.0 L** of helium at **0.980 atm and 21.0 °C**. High in the sky, the same gas fills **40.0 L** at **0.300 atm**. What is its temperature there, in °C?

- A. 111
- B. -48
- C. 225
- D. 16.1

✓ Practice 2 answer: B

$$P_1V_1 / T_1 = P_2V_2 / T_2 \rightarrow T_2 = T_1 \times P_2V_2 / (P_1V_1)$$

$$T_1 = 21.0 + 273.15 = 294.15 \text{ K} \cdot P_1V_1 = 15.68 \cdot P_2V_2 = 12.0 \text{ (L}\cdot\text{atm)}$$

$$T_2 = 294.15 \text{ K} \times \frac{0.300 \text{ atm} \times 40.0 \text{ L}}{0.980 \text{ atm} \times 16.0 \text{ L}} = 225 \text{ K} \rightarrow 225 - 273.15 = -48 \text{ }^\circ\text{C} \rightarrow \text{answer B}$$

A inverted the PV ratio: $294.15 \times 15.68 / 12.0 = 384 \text{ K} = 111 \text{ }^\circ\text{C}$, a gas that warmed as it rose. C stopped at kelvin: 225 K is the right temperature on the wrong scale. D put Celsius into the ratio: $21.0 \times 12.0 / 15.68 = 16.1 \text{ }^\circ\text{C}$.

 **Sense check:** PV fell from 15.68 to 12.0, so the kelvin temperature falls by the same factor: well below freezing, as high air is. ✓

Check yourself

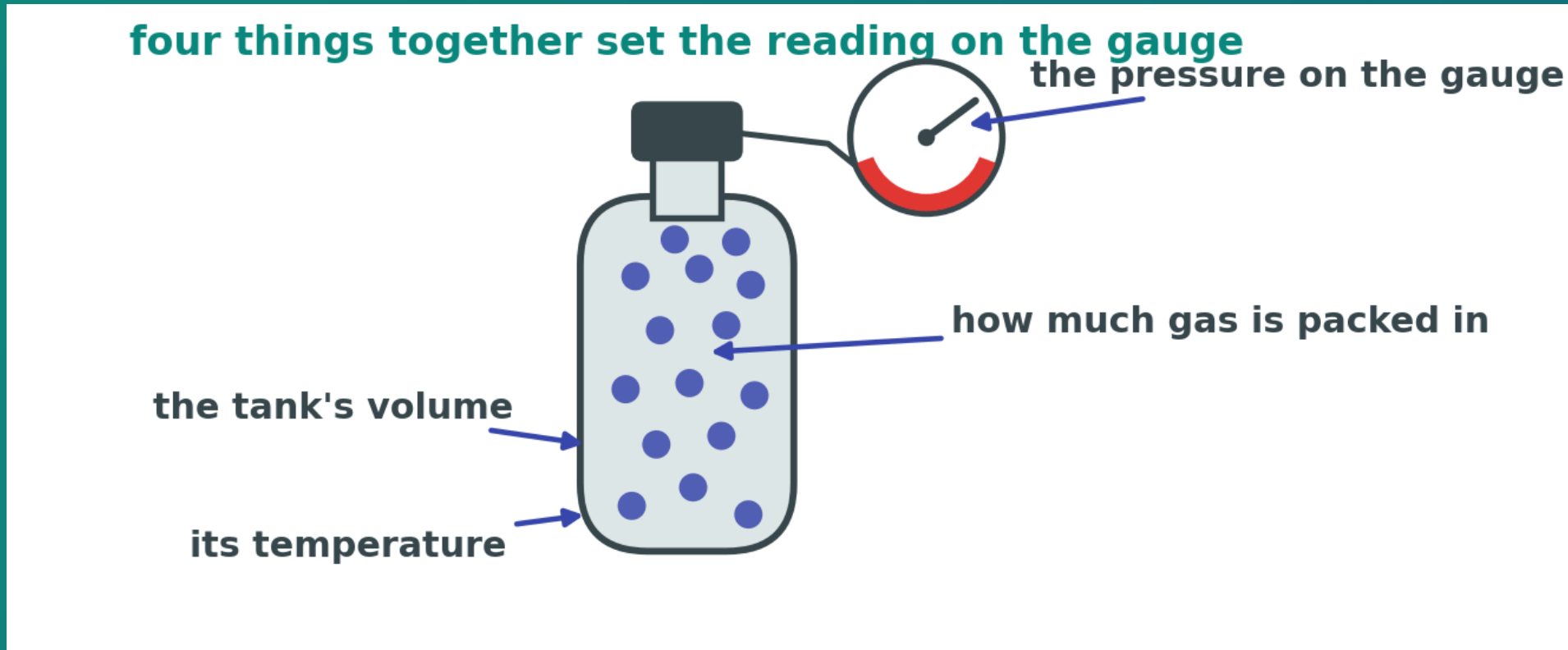
1. A gas is compressed at constant temperature. Which special case is this, and which way does the volume move relative to the pressure?
2. Write $P_1V_1/T_1 = P_2V_2/T_2$ rearranged for V_2 . When a gas is heated, which temperature goes on top of the ratio?

Add one more quantity (the amount of gas, n) and PV/T becomes a single fixed constant, R . That is the ideal gas law, $PV = nRT$: it pins down the actual size of P , V , and T , not just how they trade off.

4 • The Ideal Gas Law

Use $PV = nRT$ to find whichever of pressure, volume, amount, or temperature is unknown, with temperature in kelvins and pressure in atmospheres so the units of R cancel.

A tank's pressure comes from three things



A full scuba tank reads a high pressure. How much gas is packed in, the tank's volume, and the temperature together set that reading.

One equation ties four quantities together

$$P V = n R T$$

pressure · volume · amount of gas · temperature · bound by the constant R

A gas's pressure, volume, amount, and temperature are not independent. One equation binds all four. Know any three and the equation fixes the fourth.

The four quantities and the constant R

put every quantity in R's units: atm, L, mol, K

$$\begin{array}{ccccc} \boxed{\mathbf{P}} & \boxed{\mathbf{V}} & = & \boxed{\mathbf{n}} & \times & \boxed{\mathbf{R}} & \times & \boxed{\mathbf{T}} \\ \text{pressure} & \text{volume} & & \text{amount} & & \text{gas constant} & & \text{temperature} \\ \mathbf{atm} & \mathbf{L} & & \text{of gas} & & \mathbf{0.08206} & & \mathbf{K} \\ & & & \mathbf{mol} & & \mathbf{L \cdot atm / mol \cdot K} & & \end{array}$$

Each symbol carries a unit. R is the gas constant that links them: 0.08206 L·atm/mol·K. Its units set the units every quantity must use.

Temperature in kelvins, pressure in atm

$$T(\text{K}) = T(^{\circ}\text{C}) + 273.15 \cdot P(\text{atm}) = P(\text{mmHg}) \div 760$$

R is in L, atm, mol, K, so every quantity enters in those units

R is written in liters, atmospheres, moles, and kelvins. Every quantity must enter in those units. A temperature is always converted to kelvins; a Celsius value gives a wrong answer.

One state, with the amount included

combined gas law: $P_1V_1/T_1 = P_2V_2/T_2$ · one fixed sample, two states

ideal gas law: $PV = nRT$ · one state, with the amount n written in

The combined gas law compares a fixed sample before and after a change. The ideal gas law describes a single state and puts the amount of gas, n , directly in.



The method

1. **List the pieces:** P , V , n , T with units. Convert to kelvins and atm. Mark the unknown.
2. **Rearrange $PV = nRT$** for the unknown.
3. **Substitute $R = 0.08206$ and cancel units.**
4. **Check the units and the size.**



Worked example 1: moles in a cylinder

$$P V = n R T$$

given: 2.00 atm · 5.00 L · 25.0 °C · argon · wanted: n

A **5.00 L** cylinder holds argon at **2.00 atm** and **25.0 °C**. How many moles of argon does it hold?

List the pieces and convert the temperature to kelvins.

✓ Worked example 1: solution

$$P V = n R T$$

given: 2.00 atm · 5.00 L · 25.0 °C · wanted: n

Step 1 · List the pieces

Step 2 · Rearrange $PV = nRT$

$$P V = n R T \rightarrow n = \frac{P V}{R T}$$

Step 3 · Substitute R and cancel units

Step 4 · Check the units and the size

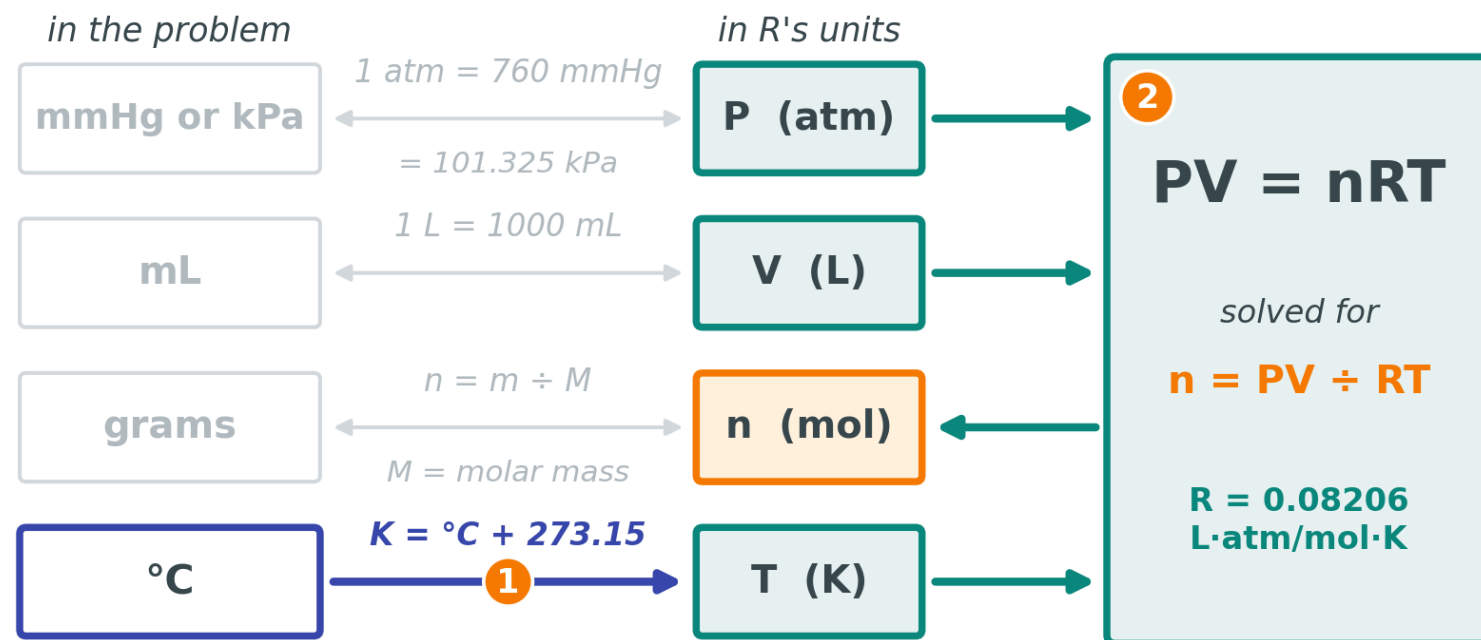
$$n = \frac{2.00 \text{ atm} \times 5.00 \text{ L}}{0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} \times 298.15 \text{ K}} = 0.409 \text{ mol}$$

 **Sense check:** atm, L, and K all cancel, leaving mol. Two atmospheres in a 5 L cylinder near room temperature comes to under half a mole of argon. ✓

✓ Worked example 1: the route on the map

$$n = PV \div RT$$

given: 2.00 atm · 5.00 L · 25.0 °C · found: n = 0.409 mol



🧠 **Sense check:** Pressure and volume were already in R's units. Only °C converted, then n came out of $PV = nRT$. ✓



Worked example 2: moles from a pressure in mmHg

$$P V = n R T$$

given: 1140 mmHg · 8.00 L · 27.0 °C · nitrogen · wanted: n

An **8.00 L** flask of nitrogen sits at **1140 mmHg** and **27.0 °C**. How many moles of nitrogen are in the flask?

The pressure is in mmHg. Convert it to atm before it enters.

✓ Worked example 2: solution

$$P V = n R T$$

given: 1140 mmHg · 8.00 L · 27.0 °C · wanted: n

Step 1 · List the pieces

$$P = 1140 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 1.50 \text{ atm}$$

Step 2 · Rearrange PV = nRT

Step 3 · Substitute R and cancel units

Step 4 · Check the units and the size

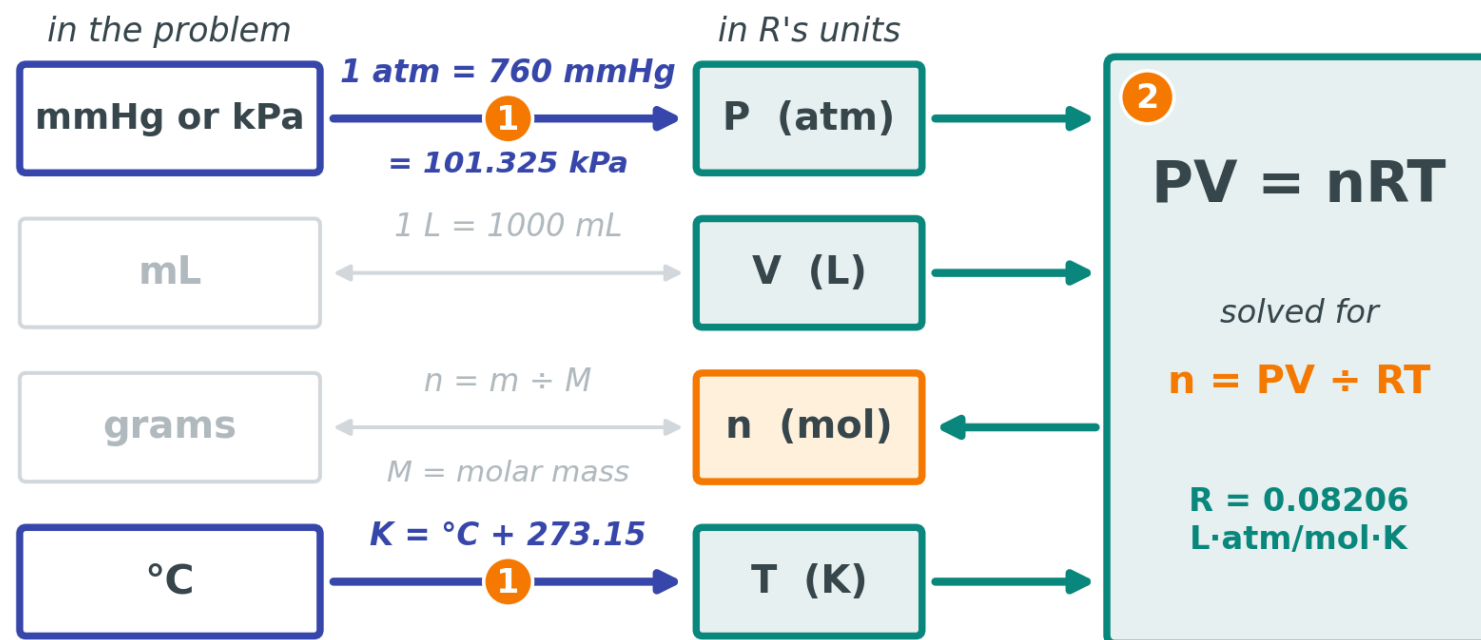
$$n = \frac{1.50 \text{ atm} \times 8.00 \text{ L}}{0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} \times 300.15 \text{ K}} = 0.487 \text{ mol}$$

 **Sense check:** The pressure entered in atm and the temperature in kelvins, so every unit but mol cancels: 0.487 mol nitrogen. ✓

✓ Worked example 2: the route on the map

$$n = PV \div RT$$

given: 1140 mmHg · 8.00 L · 27.0 °C · found: $n = 0.487$ mol



🧠 **Sense check:** Two conversions came first: mmHg to atm and °C to K. Then n came out of $PV = nRT$. ✓

🧩 Your turn: pressure in a rigid tank

$$P V = n R T$$

given: 0.500 mol · 2.00 L · 25.0 °C · wanted: P

A rigid **2.00 L** tank holds **0.500 mol** of gas at **25.0 °C**. Rearranged for pressure, $P = nRT/V$.

$$P = \frac{n R T}{V} = \frac{0.500 \text{ mol} \times \text{ } \times \text{ } \text{ K}}{2.00 \text{ L}} = \text{ } \text{ atm}$$

Fill in R and the temperature in kelvins, then compute the pressure.

$$P = \frac{0.500 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} \times 298.15 \text{ K}}{2.00 \text{ L}} = \mathbf{6.12 \text{ atm}}$$

🧠 **Sense check:** Half a mole squeezed into 2 L at room temperature pushes to about 6 atm. ✓

Where this goes wrong

$$P V = n R T$$

$$1140 \text{ mmHg (1.50 atm)} \cdot 8.00 \text{ L} \cdot 27.0 \text{ }^{\circ}\text{C (300.15 K)} \cdot \text{correct } n = 0.487 \text{ mol}$$

Leaving temperature in Celsius. Putting 27.0 in for T gives $1.50 \times 8.00 \div (0.08206 \times 27.0) = 5.42 \text{ mol}$. R is per kelvin, so $T = 27.0 + 273.15 = 300.15 \text{ K}$.

Leaving pressure in mmHg. Using 1140 gives $1140 \times 8.00 \div (0.08206 \times 300.15) = 370 \text{ mol}$, far too much gas. R uses atm: $1140 \text{ mmHg} \times (1 \text{ atm} / 760 \text{ mmHg}) = 1.50 \text{ atm}$.

Flipping the rearrangement. Writing RT/PV gives $(0.08206 \times 300.15) \div (1.50 \times 8.00) = 2.05$. Divide $PV = nRT$ by RT to isolate n : PV stays on top, $n = PV/RT$.



Practice 1

$$P V = n R T$$

given: 950. mmHg · 4.00 L · 35.0 °C · wanted: n

A **4.00 L** bulb of oxygen sits at **950. mmHg** and **35.0 °C**. How many moles of oxygen are in the bulb?

- A. 5.06
- B. 0.198
- C. 1.74
- D. 150.

✓ Practice 1 answer: B

$$P V = n R T$$

given: 950. mmHg (1.25 atm) · 4.00 L · 35.0 °C (308.15 K) · wanted: n

$$n = \frac{1.25 \text{ atm} \times 4.00 \text{ L}}{0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} \times 308.15 \text{ K}} = 0.198 \text{ mol} \rightarrow \text{answer B}$$

A flipped the rearrangement: $(0.08206 \times 308.15) \div (1.25 \times 4.00) = 5.06$. C left the temperature in Celsius: $1.25 \times 4.00 \div (0.08206 \times 35.0) = 1.74 \text{ mol}$. D left the pressure in mmHg: $950. \times 4.00 \div (0.08206 \times 308.15) = 150. \text{ mol}$.

🧠 **Sense check:** Just over 1 atm in a 4 L bulb near room temperature holds about a fifth of a mole. ✓

Molar mass and density

from $PV = nRT$, with $n = m \div M$ · $M = m R T \div (P V)$
moles = mass $\div$ molar mass · the density $d = m/V$ gives $d = P M / (R T)$

The amount n is the mass divided by the molar mass. Substituting $n = m/M$ into $PV = nRT$ solves for the molar mass, and the same swap turns density into $d = PM/RT$.

Worked example 3: identify a gas by its molar mass

$$M = m R T \div (P V)$$

given: 3.60 g · 2.00 L · 1.00 atm · 25.0 °C · wanted: M, then the gas

A **3.60 g** sample of a pure gas fills **2.00 L** at **1.00 atm** and **25.0 °C**. Candidate molar masses, in g/mol:

He	CH ₄	N ₂	O ₂	CO ₂
4.00	16.04	28.02	32.00	44.01

A common first attempt: leave the temperature at 25 °C. Find the molar mass and name the gas.

✓ Worked example 3: solution

$$M = m R T \div (P V)$$

given: 3.60 g · 2.00 L · 1.00 atm · 25.0 °C · wanted: M

A common first attempt

$$M = \frac{3.60 \text{ g} \times 0.08206 \times 25.0}{1.00 \text{ atm} \times 2.00 \text{ L}} = 3.69 \text{ g/mol} \times$$

Step 1 · List the pieces

Step 2 · Rearrange $PV = nRT$

$$P V = n R T \text{ with } n = \frac{m}{M} \rightarrow M = \frac{m R T}{P V}$$

🧠 **Sense check:** The unknown M sits inside $n = m/M$. Solving for it puts the mass on top: $M = mRT/PV$. ✓

✓ Worked example 3: the molar mass

$$M = m R T \div (P V)$$

3.60 g · 1.00 atm · 2.00 L · 298.15 K · wanted: M, then the gas

Step 3 · Substitute R and cancel units

$$M = \frac{3.60 \text{ g} \times 0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} \times 298.15 \text{ K}}{1.00 \text{ atm} \times 2.00 \text{ L}} = 44.0 \text{ g/mol}$$

He	CH ₄	N ₂	O ₂	CO ₂
4.00	16.04	28.02	32.00	44.01

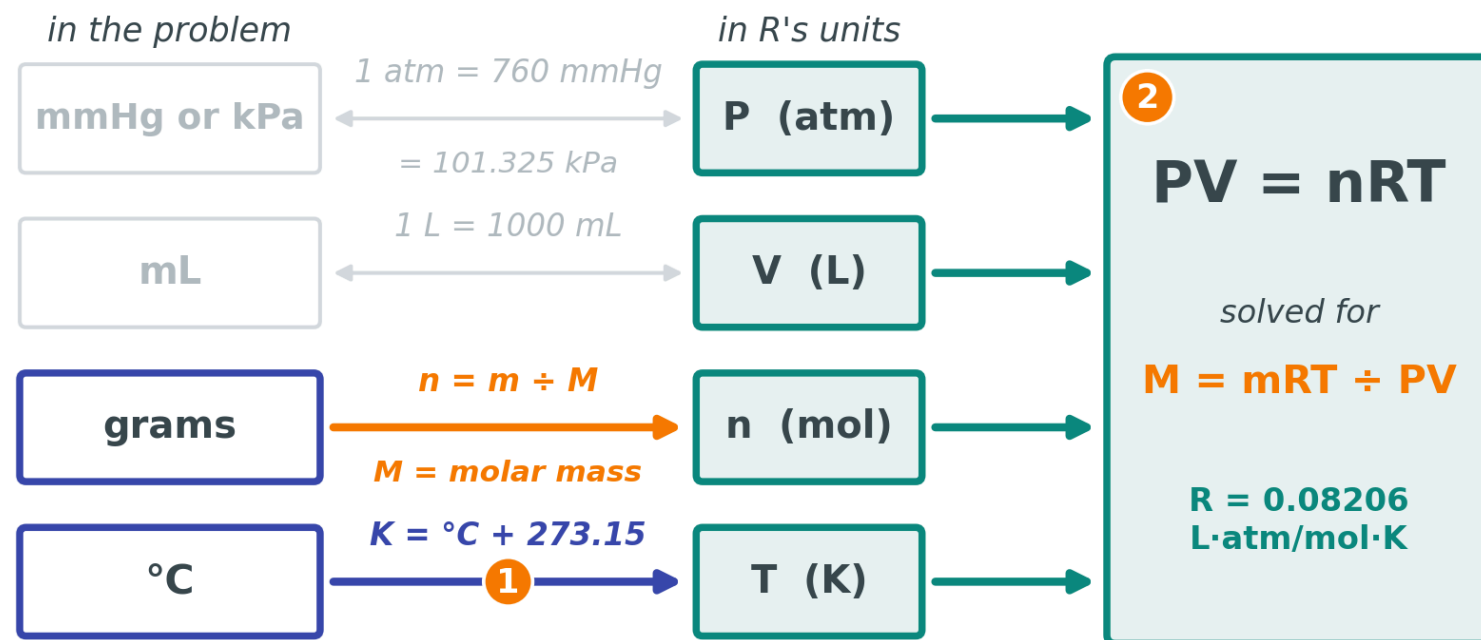
Step 4 · Check the units and the size

 **Sense check:** g on top, atm and L cancelling below, leaves g/mol. In kelvins the sample reads 44.0 g/mol, which is CO₂; left in Celsius it read 3.69 g/mol, lighter than any real gas. ✓

✓ Worked example 3: the route on the map

$$M = m R T \div (P V)$$

given: 3.60 g · 2.00 L · 1.00 atm · 25.0 °C · found: $M = 44.0 \text{ g/mol}$, CO_2



🧠 **Sense check:** Grams entered through $n = m \div M$, with M the unknown. Only °C converted. ✓

Take-home: temperature enters in kelvins

in kelvins: $T = 298.15 \text{ K} \rightarrow M = 44.0 \text{ g/mol} \cdot \text{CO}_2$
the correct molar mass ✓

in Celsius: $T = 25.0 \rightarrow M = 3.69 \text{ g/mol} \cdot \text{lighter than helium}$
impossible for a real gas ✗

R is defined per kelvin. A Celsius temperature makes every gas law answer wrong. Convert first: $T(\text{K}) = T(^{\circ}\text{C}) + 273.15$.

 Practice 2

$$d = P M \div (R T)$$

given: $d = 1.97 \text{ g/L}$ · 625 mmHg · $22.0 \text{ }^{\circ}\text{C}$ · wanted: M in g/mol

A leaking cylinder releases an unknown gas. Its density is **1.97 g/L** at **625 mmHg** and **$22.0 \text{ }^{\circ}\text{C}$** . What is the molar mass of the gas, in g/mol ?

- A. 47.7
- B. 4.32
- C. 39.2
- D. 58.0

✓ Practice 2 answer: D

$$d = P M \div (R T) \rightarrow M = d R T \div P$$

$$625 \text{ mmHg} \times (1 \text{ atm} / 760 \text{ mmHg}) = 0.822 \text{ atm} \cdot 22.0 + 273.15 = 295.15 \text{ K}$$

$$M = \frac{1.97 \text{ g/L} \times 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 295.15 \text{ K}}{0.822 \text{ atm}} = 58.0 \text{ g/mol} \rightarrow \text{answer D}$$

A stopped before dividing by P: $1.97 \times 0.08206 \times 295.15 = 47.7$, as if the pressure were 1 atm. B left the temperature in Celsius: $1.97 \times 0.08206 \times 22.0 \div 0.822 = 4.32$. C flipped the mmHg factor: $760/625 = 1.216 \text{ atm}$, and $47.7 \div 1.216 = 39.2$.

🧠 **Sense check:** 58.0 g/mol is twice air's 29 g/mol, so the gas pools low; it matches butane, C_4H_{10} , at 58.12 g/mol. ✓

Check yourself

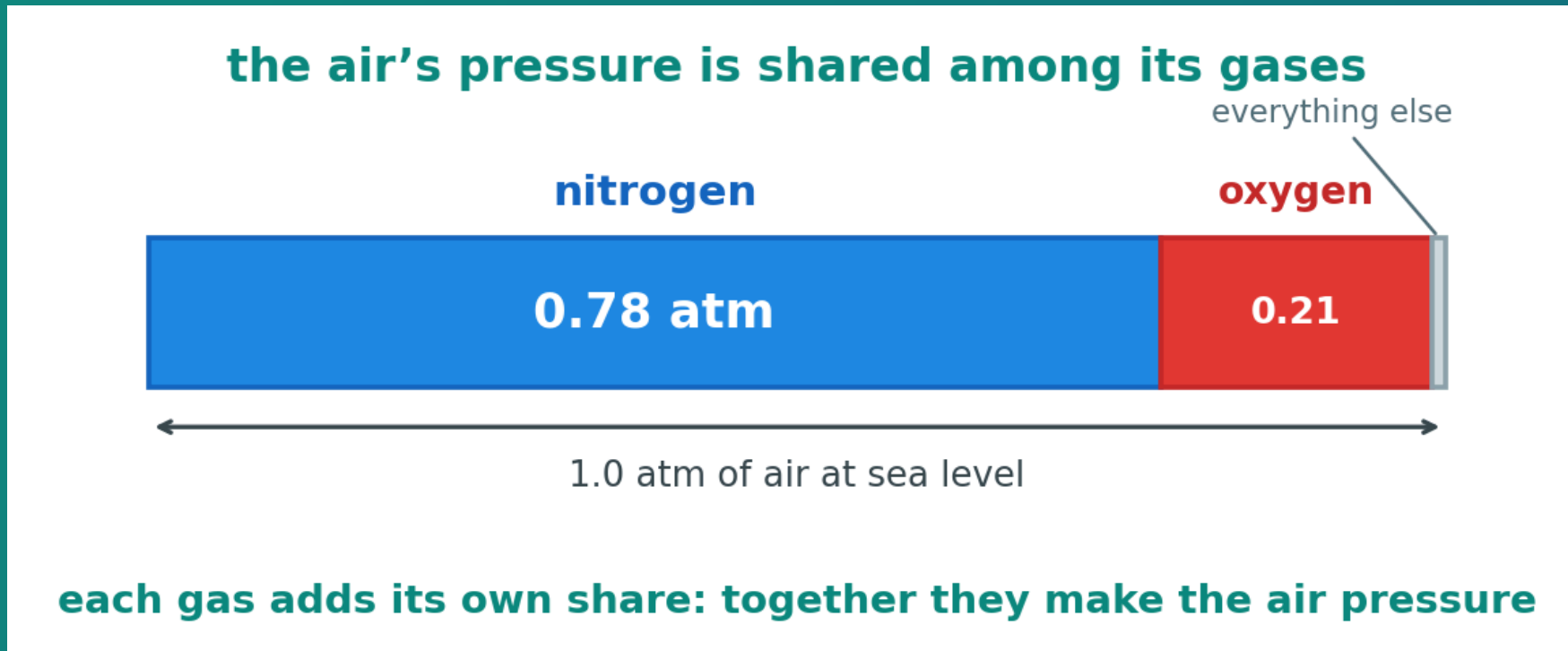
1. A gas occupies 6.00 L at 1.20 atm and 300. K. Rearrange $PV = nRT$ for n , then state which units of R cancel and which unit survives.
2. A pressure reads 745 mmHg and a temperature reads 18 °C. Convert each to the units R requires before either enters the equation.

The ideal gas law describes one gas on its own. When several gases share a container, each keeps its own partial pressure, and those pressures add. That sum is Dalton's law, and gas stoichiometry builds on the same $PV = nRT$.

5 · Dalton's Law of Partial Pressures

Find any gas's partial pressure in a mixture: the partials add to the total, and each gas's share is its mole fraction times the total.

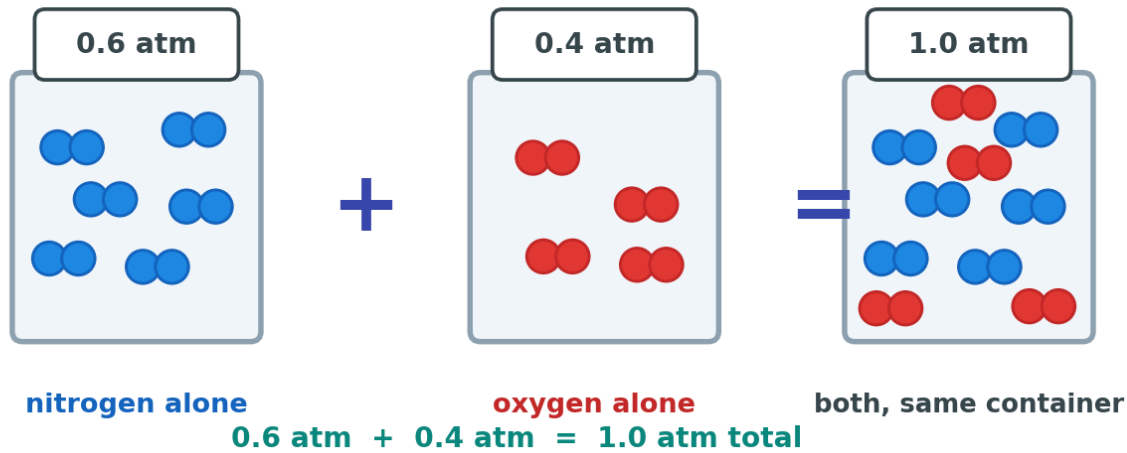
The air you breathe



Air is mostly nitrogen and oxygen. Each gas pushes with its own share of the pressure. Divers track those shares to breathe safely at depth.

Gases in a mixture act independently

each gas keeps its own pressure: the readings add



Put two gases into one container and neither disturbs the other. Each keeps the pressure it had alone. The total is the sum of these partial pressures: Dalton's law.

memory hook: the pressures just add

$P_{\text{total}} = P_A + P_B + P_C$ · each gas pushes as if it were alone in the container

What a partial pressure is

**a gas's partial pressure = the pressure it would exert alone in the whole container
same container, same temperature · the other gases do not change it**

Imagine removing every other gas and leaving one behind. The pressure that one gas would still read is its partial pressure. Adding the others back does not change it.

A gas's share of the pressure

$$P_{\text{gas}} = \text{mole fraction} \times P_{\text{total}}$$

$$\text{mole fraction} = \text{moles of that gas} \div \text{total moles of gas}$$

The molecules split the pressure the way they split the count. A gas that is one-fifth of the molecules supplies one-fifth of the total pressure.



The method

1. **List every gas** in the container.
2. **Write Dalton's law:** partials add to the total.
3. **Find the unknown:** subtract the known partials, or take its mole fraction of the total.
4. **Check:** partials rebuild the total.



Worked example 1: the missing gas

$$P_{\text{total}} = P(\text{He}) + P(\text{Ne}) + P(\text{Ar})$$

given: $P_{\text{total}} = 1.20 \text{ atm}$ · He 0.30 atm · Ne 0.35 atm · wanted: $P(\text{Ar})$

A display lamp is filled with helium, neon, and argon at a total pressure of **1.20 atm**. Helium contributes **0.30 atm** and neon **0.35 atm**. Find the partial pressure of argon.

List the gases, then write Dalton's law.

✓ Worked example 1: solution

$$P_{\text{total}} = P(\text{He}) + P(\text{Ne}) + P(\text{Ar})$$

given: $P_{\text{total}} = 1.20 \text{ atm}$ · He 0.30 atm · Ne 0.35 atm · wanted: $P(\text{Ar})$

Step 1 · List every gas

Step 2 · Write Dalton's law

Step 3 · Find the unknown

$$P_{\text{Ar}} = 1.20 \text{ atm} - 0.30 \text{ atm} - 0.35 \text{ atm} = \mathbf{0.55 \text{ atm}}$$

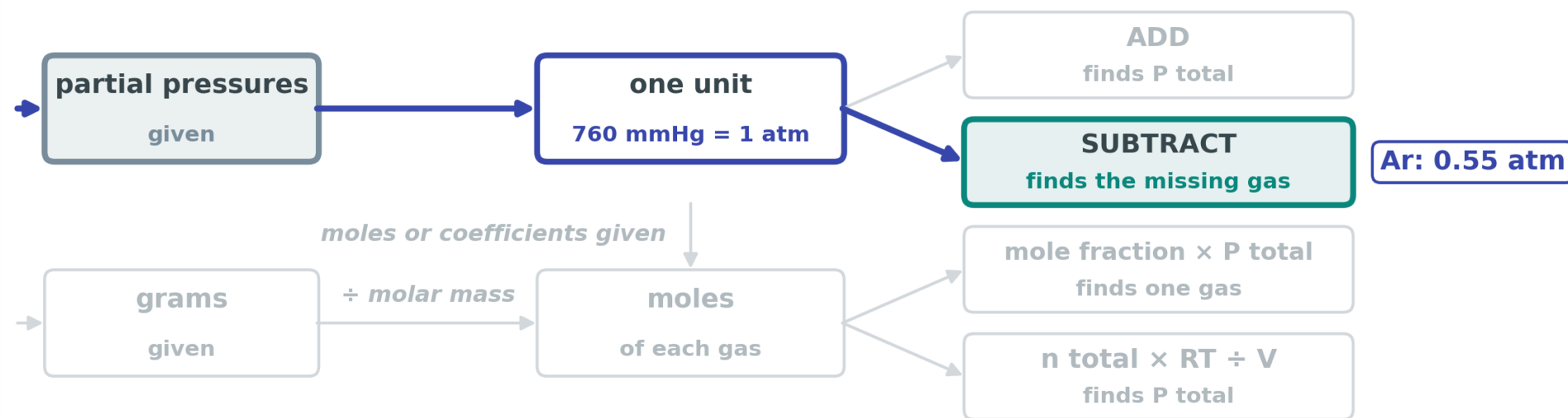
Step 4 · Check

🧠 **Sense check:** Add the three partial pressures back: $0.30 + 0.35 + 0.55 = 1.20 \text{ atm}$, the total. Argon supplies the rest of the pressure. ✓

✓ Worked example 1: the route on the map

$$P_{\text{total}} = P(\text{He}) + P(\text{Ne}) + P(\text{Ar})$$

given: $P_{\text{total}} = 1.20 \text{ atm}$ · He 0.30 atm · Ne 0.35 atm · found: $P(\text{Ar}) = 0.55 \text{ atm}$



Sense check: Every pressure was already in atm, so the unit box needed no conversion. Subtracting the known partials leaves argon. ✓



Worked example 2: a gas's share from its amount

$$P_{\text{gas}} = \text{mole fraction} \times P_{\text{total}}$$

given: 4.0 mol N₂ · 1.0 mol O₂ · P_{total} = 3.0 atm · wanted: P(O₂)

A cylinder holds **4.0 mol of nitrogen** and **1.0 mol of oxygen** at a total pressure of **3.0 atm**. Find the partial pressure of the oxygen.

Turn the amounts into oxygen's share of the molecules.

✓ Worked example 2: solution

$$P_{\text{gas}} = \text{mole fraction} \times P_{\text{total}}$$

given: 4.0 mol N₂ · 1.0 mol O₂ · P_{total} = 3.0 atm · wanted: P(O₂)

Step 1 · List every gas

Step 2 · Write Dalton's law

Step 3 · Find the unknown

$$X_{\text{O}_2} = \frac{1.0 \text{ mol O}_2}{5.0 \text{ mol total}} = 0.20$$

$$P_{\text{O}_2} = 0.20 \times 3.0 \text{ atm} = \mathbf{0.60 \text{ atm}}$$

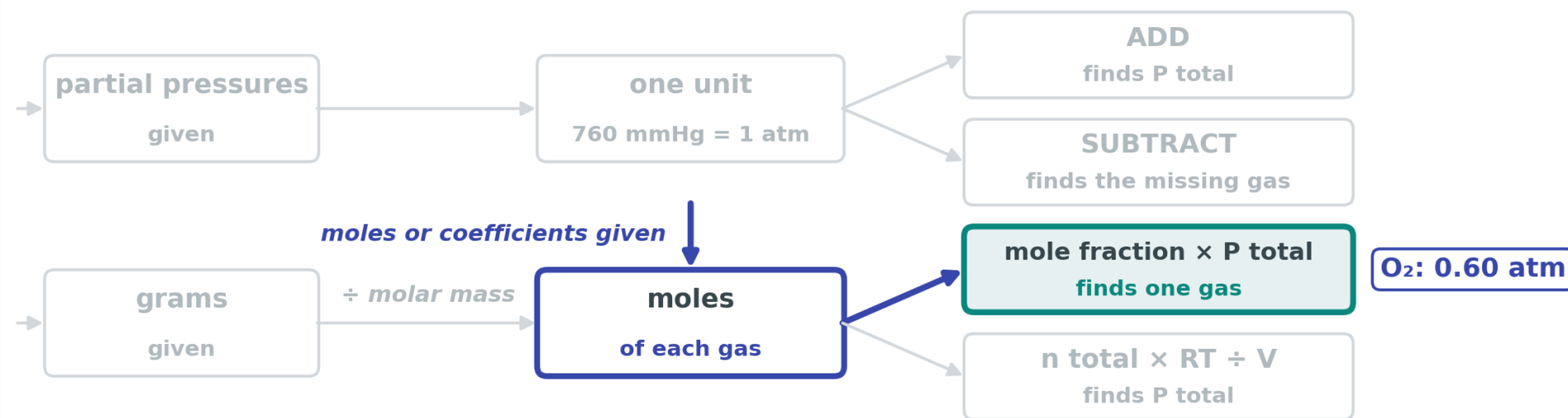
Step 4 · Check

 **Sense check:** Oxygen is one-fifth of the molecules, so it carries one-fifth of 3.0 atm: 0.60 atm. Nitrogen takes the other 2.40 atm, and 0.60 + 2.40 = 3.00. Flipping the fraction to 5.0 ÷ 1.0 would give 15 atm, more than the whole mixture. ✓

✓ Worked example 2: the route on the map

$$P_{\text{gas}} = \text{mole fraction} \times P_{\text{total}}$$

given: 4.0 mol N₂ · 1.0 mol O₂ · P_{total} = 3.0 atm · found: P(O₂) = 0.60 atm



Sense check: Moles were given, so the route starts at the moles box and leaves through mole fraction × P total. No grams, no unit change. ✓

Your turn: helium's partial pressure

$$P_{\text{gas}} = \text{mole fraction} \times P_{\text{total}}$$

given: 3.0 mol He · 1.0 mol Ne · $P_{\text{total}} = 2.0 \text{ atm}$ · wanted: $P(\text{He})$

A lab mixes **3.0 mol of helium** with **1.0 mol of neon**; the gauge reads **2.0 atm**. Fill the mole fraction from the amounts, then find helium's partial pressure.

$$P_{\text{He}} = \frac{\text{_____ mol He}}{\text{_____ mol total}} \times 2.0 \text{ atm} = \text{_____ atm}$$

$$P_{\text{He}} = \frac{3.0 \text{ mol He}}{4.0 \text{ mol total}} \times 2.0 \text{ atm} = \mathbf{1.50 \text{ atm}}$$

 **Sense check:** Helium is three-fourths of the molecules, so it carries three-fourths of the 2.0 atm. Neon takes the rest: $0.25 \times 2.0 = 0.50 \text{ atm}$, and $1.50 + 0.50 = 2.00$. ✓

⚠ Where this goes wrong

He + Ne + Ar in one container

$$P_{\text{total}} = 1.20 \text{ atm} \cdot \text{He } 0.30 \text{ atm} \cdot \text{Ne } 0.35 \text{ atm} \cdot \text{correct } P(\text{Ar}) = 0.55 \text{ atm}$$

Subtracting only one gas. $1.20 - 0.30 = 0.90 \text{ atm}$ still holds neon's 0.35 inside. Every gas takes a share of the total, so subtract both known ones: $1.20 - 0.30 - 0.35 = 0.55 \text{ atm}$.

Splitting the total into equal shares. $1.20 \div 3 = 0.40 \text{ atm}$ assumes the three gases are present in equal amounts. They are not. Each gas keeps its own pressure; subtract the known ones from the total.

Adding the known pressures. $0.30 + 0.35 = 0.65 \text{ atm}$ is the combined push of helium and neon, not argon's. Argon supplies the rest: $1.20 - 0.65 = 0.55 \text{ atm}$.



Practice 1

$$P_{\text{total}} = P(\text{N}_2) + P(\text{O}_2) + P(\text{CO}_2)$$

given: $P_{\text{total}} = 1.30 \text{ atm}$ · N_2 0.50 atm · O_2 0.20 atm · wanted: $P(\text{CO}_2)$

A sealed flask holds nitrogen, oxygen, and carbon dioxide at a total pressure of **1.30 atm**. Nitrogen contributes **0.50 atm** and oxygen **0.20 atm**. What is the partial pressure of carbon dioxide, in atm?

- A. 0.60
- B. 0.80
- C. 1.10
- D. 0.70

✓ Practice 1 answer: A

$$P_{\text{total}} = P(\text{N}_2) + P(\text{O}_2) + P(\text{CO}_2)$$

$$P_{\text{total}} = 1.30 \text{ atm} \cdot \text{N}_2 \text{ } 0.50 \text{ atm} \cdot \text{O}_2 \text{ } 0.20 \text{ atm}$$

$$P_{\text{CO}_2} = 1.30 \text{ atm} - 0.50 \text{ atm} - 0.20 \text{ atm} = \mathbf{0.60 \text{ atm} \rightarrow \text{answer A}}$$

B removed only nitrogen: $1.30 - 0.50 = 0.80 \text{ atm}$, leaving oxygen's share inside. C removed only oxygen: $1.30 - 0.20 = 1.10 \text{ atm}$, leaving nitrogen's share inside. D added the two known pressures: $0.50 + 0.20 = 0.70 \text{ atm}$, the combined push of nitrogen and oxygen.

 **Sense check:** Carbon dioxide supplies whatever the other two do not. Add all three back: $0.50 + 0.20 + 0.60 = 1.30 \text{ atm}$. ✓



Worked example 3: a reaction sets the shares



given: the products reach $P_{\text{total}} = 866.0 \text{ mmHg}$ · wanted: $P(\text{N}_2)$ and $P(\text{H}_2)$

Ammonia decomposes completely inside a sealed vessel. The product gases reach a total pressure of **866.0 mmHg**. Find the partial pressure of each product.

Read the amounts from the coefficients.

✓ Worked example 3: solution



given: the products reach $P_{\text{total}} = 866.0 \text{ mmHg}$ · wanted: $P(\text{N}_2)$ and $P(\text{H}_2)$

Step 1 · List every gas

Step 2 · Write Dalton's law

Step 3 · Find the unknown

$$P_{\text{N}_2} = \frac{1 \text{ mol N}_2}{4 \text{ mol product}} \times 866.0 \text{ mmHg} = \mathbf{216.5 \text{ mmHg}}$$

$$P_{\text{H}_2} = \frac{3 \text{ mol H}_2}{4 \text{ mol product}} \times 866.0 \text{ mmHg} = \mathbf{649.5 \text{ mmHg}}$$

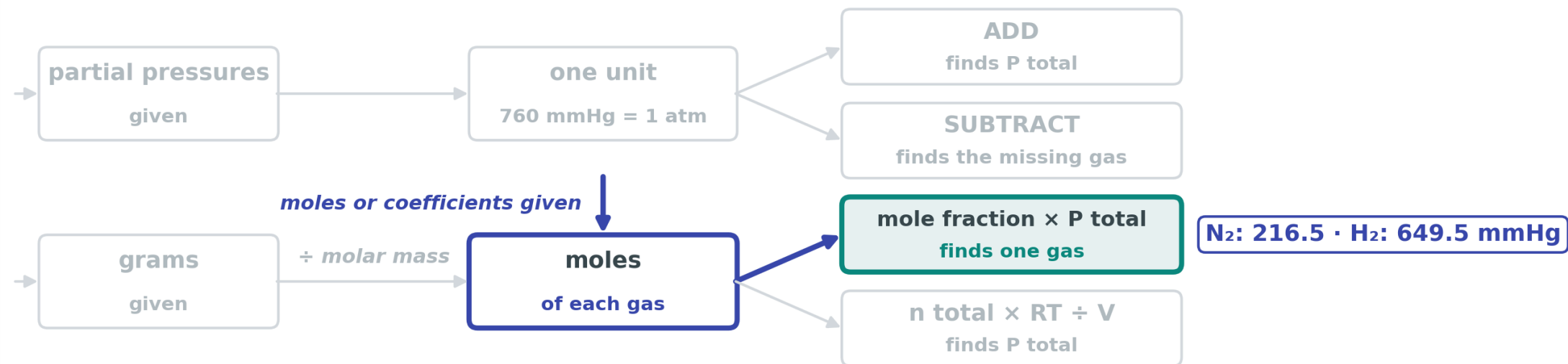
Step 4 · Check

 **Sense check:** $216.5 + 649.5 = 866.0 \text{ mmHg}$, the total. Hydrogen is three-fourths of the molecules and carries three-fourths of the pressure. Splitting the total in half, 433.0 mmHg each, ignores the 1 : 3 ratio. ✓

✓ Worked example 3: the route on the map



coefficients: 1 mol N₂ : 3 mol H₂ · found: N₂ 216.5 mmHg · H₂ 649.5 mmHg



Sense check: The coefficients stand in for moles. The route is the moles lane, walked once for each gas. ✓



Practice 2

$$P_{\text{gas}} = \text{mole fraction} \times P_{\text{total}}$$

given: 36.0 g Ar · 6.00 g He · $P_{\text{total}} = 1.50 \text{ atm}$ · wanted: $P(\text{Ar})$

A sealed flask of welding shield gas holds **36.0 g of argon** and **6.00 g of helium** at a total pressure of **1.50 atm**. What is the partial pressure of argon, in atm?

- A. 1.29
- B. 4.00
- C. 0.563
- D. 0.901

✓ Practice 2 answer: C

$$P_{\text{gas}} = \text{mole fraction} \times P_{\text{total}}$$

$$\text{Ar: } 36.0 \text{ g} \div 39.95 \text{ g/mol} = 0.901 \text{ mol} \cdot \text{He: } 6.00 \text{ g} \div 4.003 \text{ g/mol} = 1.499 \text{ mol} \cdot 2.400 \text{ mol total}$$

$$X_{\text{Ar}} = \frac{0.901 \text{ mol Ar}}{2.400 \text{ mol total}} = 0.375 \cdot P_{\text{Ar}} = 0.375 \times 1.50 \text{ atm} = \mathbf{0.563 \text{ atm} \rightarrow \text{answer C}}$$

A used the mass fraction and skipped the molar masses: $36.0 \div 42.0 \times 1.50 = 1.29 \text{ atm}$. Pressure counts particles, not grams. B flipped the mole fraction: $2.400 \div 0.901 \times 1.50 = 4.00 \text{ atm}$, more than the whole mixture. D stopped at moles: 0.901 mol of argon is an amount, not a pressure.

 **Sense check:** Argon is most of the mass but only 37.5% of the atoms. Helium carries the rest: $1.50 - 0.563 = 0.937 \text{ atm}$. ✓

Check yourself

1. A flask holds N_2 , O_2 , and CO_2 at a total of 1.10 atm. N_2 contributes 0.70 atm and O_2 0.25 atm. Write CO_2 's partial pressure and show the three rebuild the total.
2. A cylinder holds 6.0 mol of H_2 and 2.0 mol of N_2 at a total pressure of 4.0 atm. Find hydrogen's mole fraction, then its partial pressure.

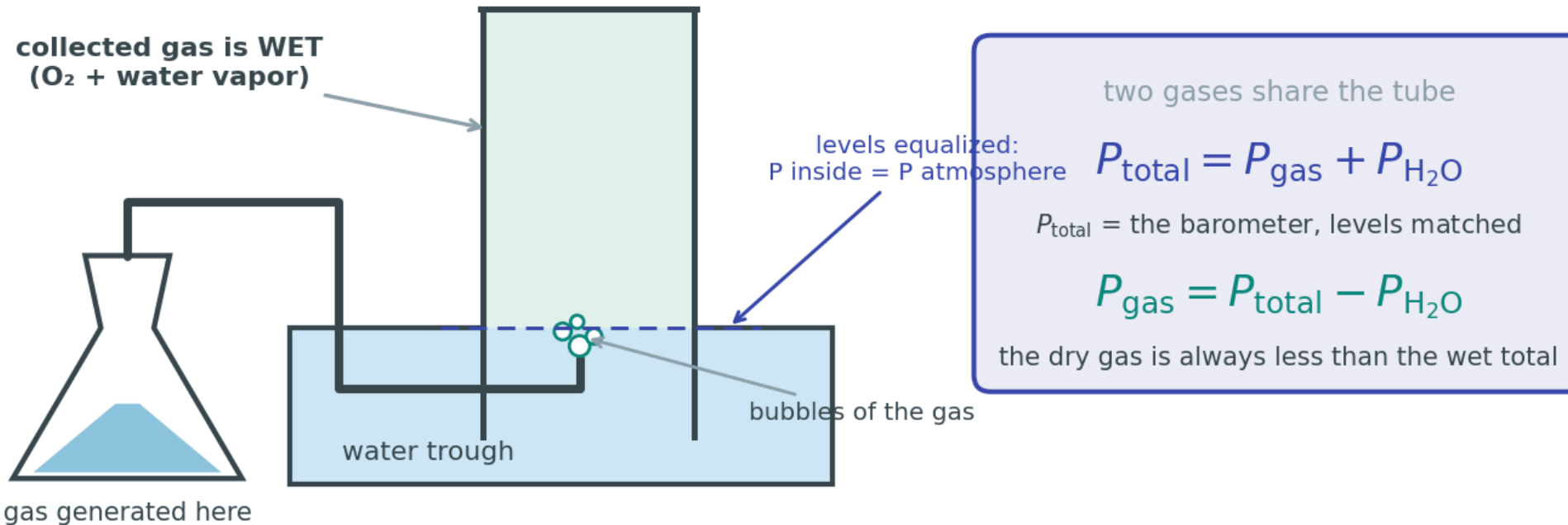
A gas bubbled up through water is a mixture too: the gas plus water vapor. Subtracting the vapor's partial pressure leaves the dry gas, and that pressure feeds the ideal gas law.

6 · Collecting a Gas Over Water

For a gas collected by water displacement, subtract the water-vapor pressure from the leveled total to get the dry-gas pressure, then feed that dry pressure, never the total, into the ideal gas law.

The gas comes out wet

Collecting a gas over water: the pressure bookkeeping



Oxygen from decomposing hydrogen peroxide bubbles up through water into an inverted tube. The gas caught this way is never pure. It carries water vapor with it.

The collected sample is a mixture

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

the tube holds the gas you want PLUS water vapor: the sample is "wet"

Above liquid water, the space is always damp: some water evaporates and pushes with its own pressure. So the tube holds two gases, and by Dalton's law their partial pressures add to the total.

memory hook: over water, subtract the water first

$P_{\text{gas}} = P_{\text{total}} - P_{\text{H}_2\text{O}}$ · the vapor is already inside the reading, never add it

Leveling makes the total easy to read

raise or lower the tube until the water is level inside and out

then $P_{\text{inside}} = P_{\text{atmosphere}}$, so $P_{\text{total}} = \text{the barometer reading}$

Unequal water levels add or subtract a pressure of their own. Match the level inside to the trough outside, and the wet gas pushes at exactly atmospheric pressure.

Water vapor's pressure depends only on temperature

$P_{\text{H}_2\text{O}}$ is fixed by temperature alone: look it up in a table

15 °C → 12.8 torr · 20 °C → 17.5 torr · 25 °C → 23.8 torr · 30 °C → 31.8 torr

Water's vapor pressure depends on temperature alone, not on the gas, the tube, or the barometer. Warmer water evaporates more. Read the value from the table at the water's temperature.



The method

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

leveled tube: $P_{\text{total}} = \text{barometer} \cdot P_{\text{H}_2\text{O}}$ from the table $\cdot$ dry pressure into $PV = nRT$

1. **Level and read.** Levels matched, the barometer gives P_{total} .
2. **Look up $P_{\text{H}_2\text{O}}$** at the water's temperature.
3. **Subtract** the vapor for the dry gas.
4. **Use the dry pressure** in $PV = nRT$, in atm and kelvin.



Worked example 1: oxygen over water

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

given: O₂ over water · 25 °C · barometer 755.0 torr · wanted: P(O₂)

Oxygen from decomposing hydrogen peroxide is collected over water at **25 °C**. With the tube leveled, the barometer reads **755.0 torr**. Find the pressure of the dry oxygen.

Level, look up, subtract.

✓ Worked example 1: solution

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

given: O₂ over water · 25 °C · barometer 755.0 torr · wanted: P(O₂)

Step 1 · Level and read

Step 2 · Look up P_{H₂O}

Step 3 · Subtract

$$P_{\text{O}_2} = 755.0 \text{ torr} - 23.8 \text{ torr} = \mathbf{731.2 \text{ torr}}$$

 **Sense check:** The dry oxygen sits a little below the 755.0 torr wet total, as a part must. Adding instead gives 778.8 torr, a part larger than the whole. ✓

✓ Worked example 1: the route on the map

$$P_{O_2} = 755.0 \text{ torr} - 23.8 \text{ torr} = 731.2 \text{ torr}$$

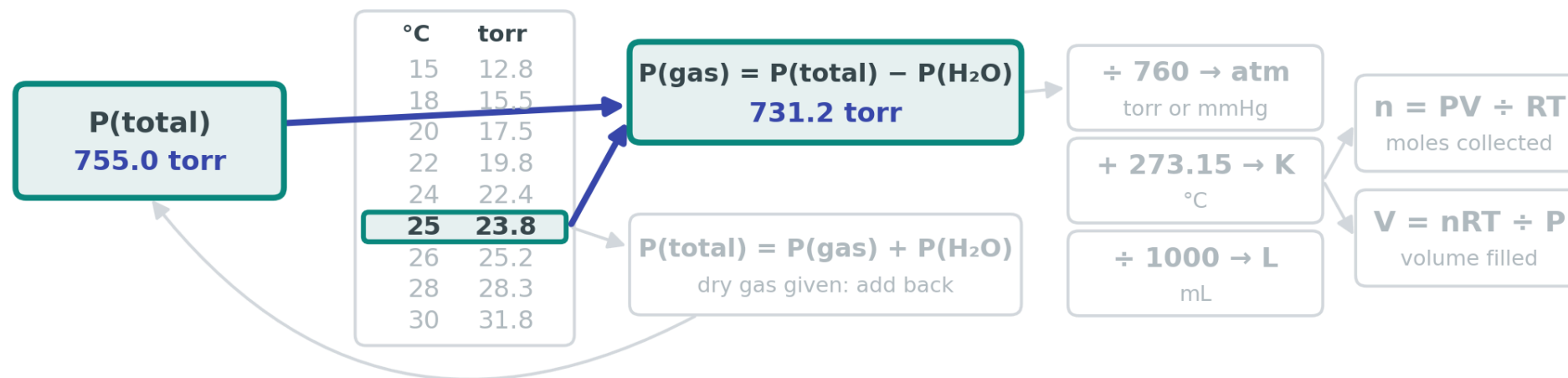
given: barometer 755.0 torr · 25 °C row: 23.8 torr · found: dry O₂ 731.2 torr

1 · Level and read

2 · Look up P(H₂O)

3 · Subtract

4 · Use the dry pressure



Sense check: Three moves: read the leveled total, look up the 25 °C row, subtract. Step 4 stays grey: the question stopped at the dry pressure. ✓



Worked example 2: hydrogen over water, to moles

$$P_{\text{gas}} = P_{\text{total}} - P_{\text{H}_2\text{O}}, \text{ then } n = P_{\text{gas}}V / RT$$

given: H_2 over water · 24 °C · barometer 762 torr · $V = 0.250 \text{ L}$ · wanted: mol H_2

A student collects hydrogen over water at **24 °C**. The barometer reads **762 torr** and the leveled tube holds **0.250 L**. How many moles of H_2 were collected?

Find the dry pressure first, then let it drive the ideal gas law.

✓ Worked example 2: the dry pressure

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

given: H₂ over water · 24 °C · barometer 762 torr · V = 0.250 L · wanted: mol H₂

Step 1 · Level and read

Step 2 · Look up P_{H₂O}

Step 3 · Subtract

$$P_{\text{H}_2} = 762 \text{ torr} - 22.4 \text{ torr} = \mathbf{739.6 \text{ torr}}$$

🧠 **Sense check:** The dry hydrogen sits just below the 762 torr wet total, as a part must. ✓

✓ Worked example 2: the moles

$$n = P_{\text{gas}} V / (RT)$$

$P_{\text{H}_2} = 739.6 \text{ torr} \cdot V = 0.250 \text{ L} \cdot T = 24 \text{ }^\circ\text{C} \cdot R = 0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$ needs atm and kelvin

Step 4 • Use the dry pressure

$$739.6 \text{ ~~torr~~} \times \frac{1 \text{ atm}}{760 \text{ ~~torr~~}} = \mathbf{0.973 \text{ atm}} \cdot T = 24 + 273.15 = \mathbf{297.15 \text{ K}}$$

$$n = \frac{0.973 \text{ atm} \times 0.250 \text{ L}}{0.08206 \times 297.15 \text{ K}} = \mathbf{0.00998 \text{ mol}}$$

 **Sense check:** Use the wet total 762 torr instead and n comes out 0.0103 mol, about 3% too high. The vapor was never hydrogen. ✓

✓ Worked example 2: the route on the map

$$n = P_{\text{gas}}V / (RT)$$

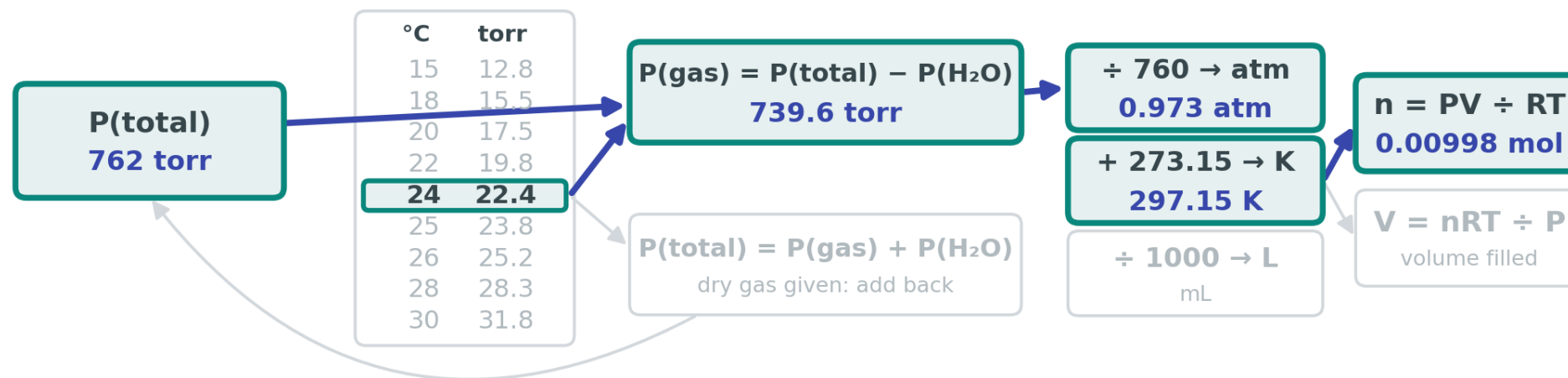
given: barometer 762 torr · 24 °C row: 22.4 torr · V = 0.250 L · found: 0.00998 mol H₂

1 · Level and read

2 · Look up P(H₂O)

3 · Subtract

4 · Use the dry pressure



🧠 **Sense check:** All four steps are lit. The dry 739.6 torr, not the wet 762 torr, is the pressure that reaches $n = PV \div RT$.



Your turn: nitrogen over water

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

given: N₂ over water · 22 °C · barometer 750. torr · P_{H₂O} = 19.8 torr · wanted: P(N₂) in atm

Nitrogen is collected over water at **22 °C** in a leveled tube; the barometer reads **750. torr** and water vapor is **19.8 torr**. Fill the blanks, then convert to atm.

$$P_{\text{N}_2} = (750. \text{ torr} - \text{ ______ torr}) \times \frac{1 \text{ atm}}{760 \text{ torr}} = \text{ ______ atm}$$

$$P_{\text{N}_2} = (750. \text{ torr} - 19.8 \text{ torr}) \times \frac{1 \text{ atm}}{760 \text{ torr}} = 730.2 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = \mathbf{0.961 \text{ atm}}$$

 **Sense check:** The subtraction leaves 730.2 torr of dry nitrogen, just under the 750. torr wet total. PV = nRT takes 0.961 atm, not 750 ÷ 760 = 0.987 atm. ✓

⚠ Where this goes wrong

Using P_{total} for the gas. Feeding the full 762 torr into $PV = nRT$ counts the water vapor as if it were H_2 (0.0103 vs 0.00998 mol). Subtract first: $P_{\text{gas}} = P_{\text{total}} - P_{\text{H}_2\text{O}}$.

Adding the vapor. $762 + 22.4 = 784.4$ torr makes one gas push harder than the whole wet mixture. The vapor is already inside the reading: subtract it, never add.

Skipping the conversions. $R = 0.08206$ wants atm and kelvin: $739.6 \text{ torr} \div 760 = 0.973 \text{ atm}$, $24^\circ\text{C} + 273.15 = 297.15 \text{ K}$. And only a leveled tube makes P_{total} equal the barometer.



Practice 1

$$P_{\text{atm}} = P_{\text{CH}_4} + P_{\text{H}_2\text{O}}$$

given: CH₄ over water · 25 °C · levels matched · P_{CH₄} = 727.4 torr · table: 20 °C → 17.5 torr, 25 °C → 23.8 torr · wanted: P_{atm}

Methane bubbles into an inverted tube standing in **25 °C** water, and the water levels inside and out are matched. The methane alone exerts **727.4 torr**. What is the atmospheric pressure in the room, in torr?

- A. 751.2
- B. 703.6
- C. 744.9
- D. 727.4

✓ Practice 1 answer: A

$$P_{\text{atm}} = P_{\text{CH}_4} + P_{\text{H}_2\text{O}}$$

levels matched: the wet gas balances the atmosphere · CH₄ 727.4 torr · water vapor 23.8 torr at 25 °C

$$P_{\text{atm}} = 727.4 \text{ torr} + 23.8 \text{ torr} = \mathbf{751.2 \text{ torr} \rightarrow \text{answer A}}$$

B subtracted the vapor out of habit: $727.4 - 23.8 = 703.6$ torr, but the dry pressure is already the part; the whole is larger. C read the 20 °C row: $727.4 + 17.5 = 744.9$ torr, the wrong temperature. D skipped the vapor: 727.4 torr is the methane alone, not the whole wet mixture that balances the atmosphere.

 **Sense check:** The atmosphere balances the whole wet sample, so it must sit above the methane's own 727.4 torr, by exactly the vapor. ✓

 Practice 2

$$V = n R T / P_{\text{gas}}$$

given: 0.0150 mol O₂ over water · 20.0 °C · barometer 748 torr · P_{H₂O} = 17.5 torr · wanted: V in L

A leveled tube collects **0.0150 mol** of oxygen over water at **20.0 °C**. The barometer reads **748 torr**, and water vapor at 20.0 °C is **17.5 torr**. What volume, in L, does the gas occupy?

- A. 0.367
- B. 4.94×10^{-4}
- C. 0.0256
- D. 0.375

✓ Practice 2 answer: D

$$V = n R T / P_{\text{gas}}$$

barometer 748 torr · water vapor 17.5 torr at 20.0 °C · $n = 0.0150 \text{ mol}$ · $T = 20.0 + 273.15 = 293.15 \text{ K}$

$$P_{\text{O}_2} = 748 \text{ torr} - 17.5 \text{ torr} = \mathbf{730.5 \text{ torr}} = \mathbf{0.961 \text{ atm}}$$

$$V = \frac{0.0150 \cancel{\text{ mol}} \times 0.08206 \text{ L} \cdot \cancel{\text{ atm}} / \cancel{\text{ mol}} \cdot \cancel{\text{ K}} \times 293.15 \cancel{\text{ K}}}{0.961 \cancel{\text{ atm}}} = \mathbf{0.375 \text{ L} \rightarrow \text{answer D}}$$

A used the wet total: $748 \div 760 = 0.984 \text{ atm}$ in the denominator gives 0.367 L, with the vapor's share counted as oxygen. B left the pressure in torr: dividing by 730.5 gives $4.94 \times 10^{-4} \text{ L}$. C left the temperature in Celsius: 0.08206×20.0 on top gives 0.0256 L.

 **Sense check:** Near room temperature and 1 atm a mole fills about 24 L, so 0.0150 mol fills about 0.36 L. The dry pressure sits below 1 atm, so the volume runs a little larger: 0.375 L. ✓

Check yourself

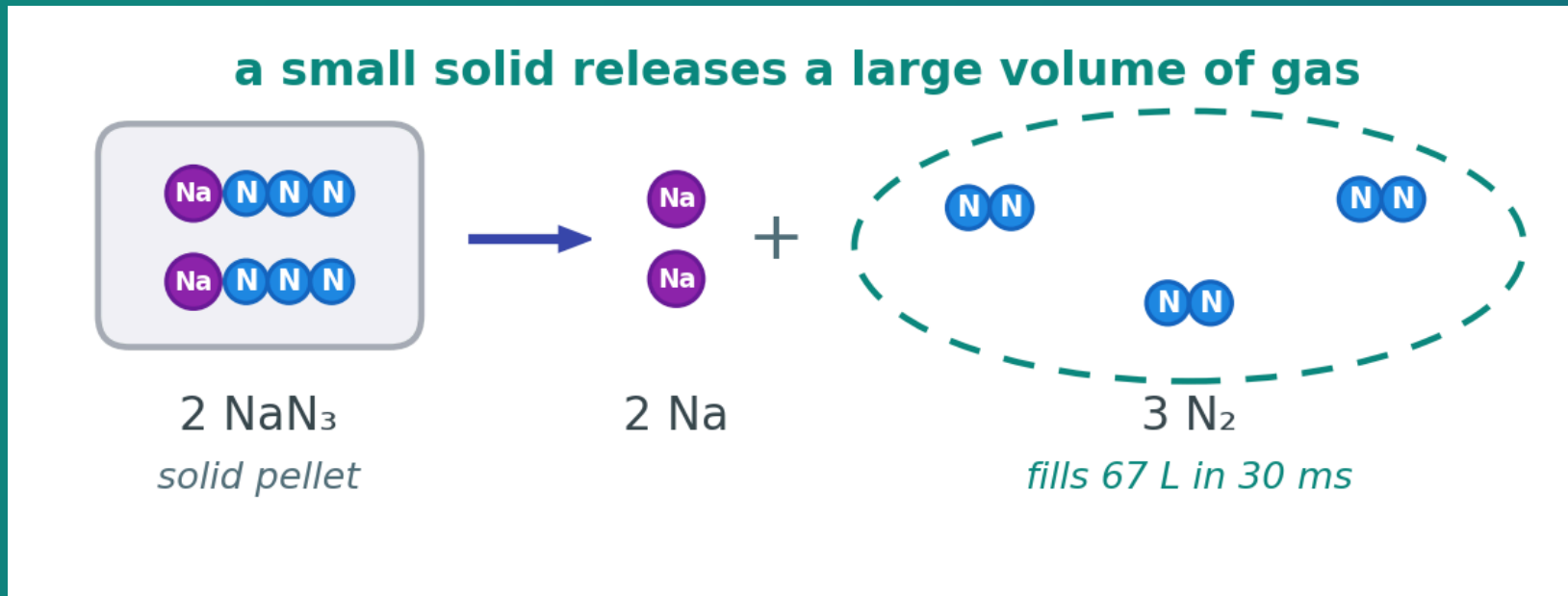
1. Hydrogen is collected over water at 30 °C in a leveled tube. The barometer reads 770 torr and water vapor at 30 °C is 31.8 torr. Find the dry H₂ pressure, and state which pressure you would put into $PV = nRT$.
2. Why does using the barometer reading directly always overcount the moles of gas collected?

Subtracting the water vapor turns a wet reading into the dry partial pressure. That pressure feeds straight into $PV = nRT$, and from there into a reaction's mole ratios, in gas stoichiometry.

7 • Gas Stoichiometry

Bring a gas volume into the mole bridge (as moles at STP through 22.4 L/mol, or off STP through $PV = nRT$), then cross the mole ratio to reach grams or a second gas volume.

An airbag, in milliseconds



A crash sensor fires, and a solid pellet decomposes. In about 30 milliseconds it releases enough gas to fill a 67-liter airbag. Stoichiometry sets the amount.

A gas volume counts moles

1 mol of any gas = 22.4 L at STP

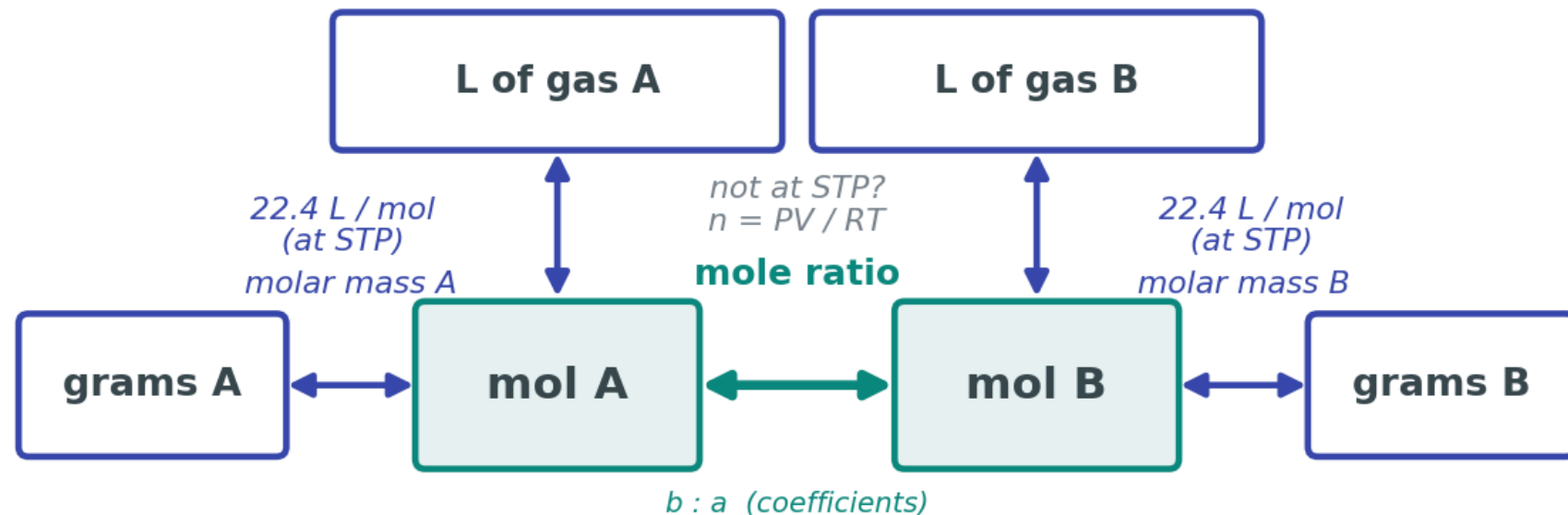
STP: 0 °C (273.15 K), 1 atm · every gas, same volume per mole

At STP one mole of any gas fills 22.4 L. Measure a volume, and this factor converts it to moles; the liters cancel:

$$11.2 \text{ L gas} \times \frac{1 \text{ mol gas}}{22.4 \text{ L gas}} = \mathbf{0.500 \text{ mol gas}}$$

Once it is moles, stoichiometry proceeds exactly as before.

Volume joins the map



a gas volume counts moles — 22.4 L/mol at STP, else $PV = nRT$

A gas volume enters through the molar volume, exactly where grams enter through molar mass. Every route still crosses the mole bridge, and the mole ratio still switches substances.

When the gas is not at STP

$$n = PV / RT$$

$R = 0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$ · T in kelvin, not °C

The 22.4 L/mol shortcut holds only at STP: 0 °C and 1 atm. At any other conditions the ideal-gas law counts the moles directly.

at STP: $0.08206 \times 273.15 = 22.4 \text{ L per mole}$

the molar volume is $PV = nRT$ evaluated at 0 °C, 1 atm



The method

1. **To moles:** molar mass for grams; 22.4 L/mol for a gas volume.
2. **Moles → moles:** cross the mole ratio. Nothing else switches substance.
3. **To the wanted unit:** the same factors, in reverse.

Off STP, $n = PV/RT$ replaces it.



Worked example 1



given: 33.6 L H₂ at STP · wanted: g NH₃

Ammonia is made from hydrogen and nitrogen. What mass of NH₃ forms when **33.6 L of H₂**, measured at STP, reacts with excess N₂? (NH₃ 17.03 g/mol)

Write the route first: L H₂ → mol H₂ → mol NH₃ → g NH₃.

✓ Worked example 1: solution



given: 33.6 L H₂ at STP · wanted: g NH₃

Step 1 · To moles

$$33.6 \cancel{\text{L H}_2} \times \frac{1 \text{ mol H}_2}{22.4 \cancel{\text{L H}_2}} = 1.50 \text{ mol H}_2$$

Step 2 · Moles → moles

Step 3 · To the wanted unit

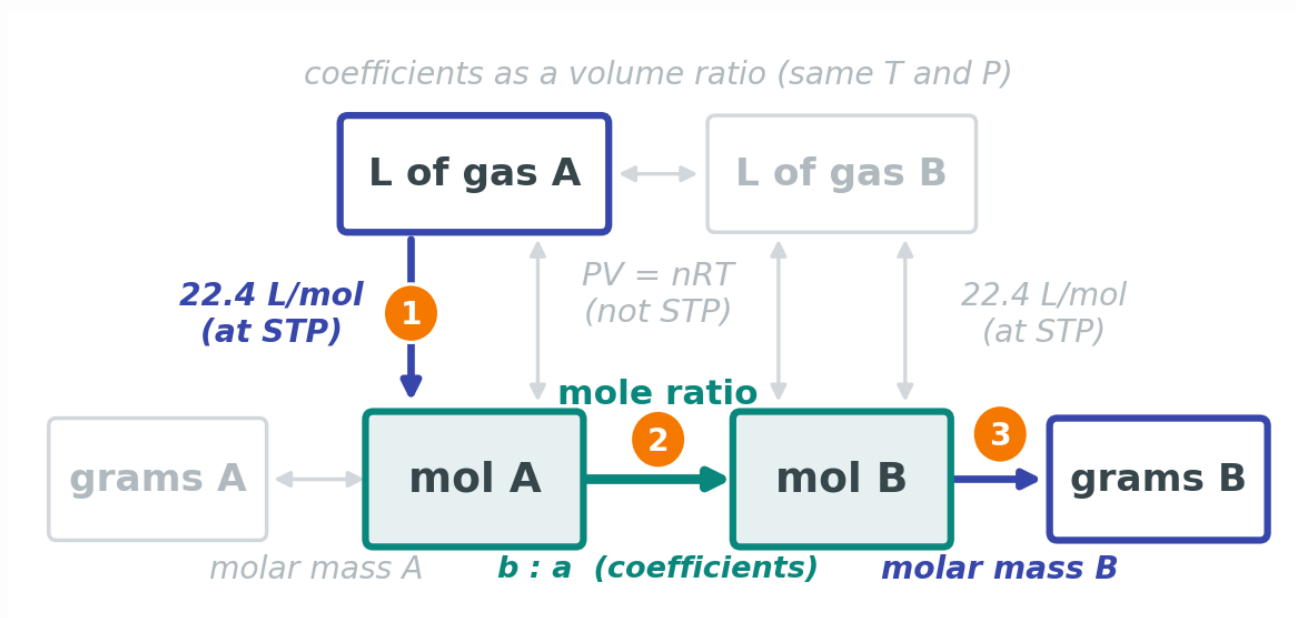
$$33.6 \cancel{\text{L H}_2} \times \frac{1 \cancel{\text{mol H}_2}}{22.4 \cancel{\text{L H}_2}} \times \frac{2 \cancel{\text{mol NH}_3}}{3 \cancel{\text{mol H}_2}} \times \frac{17.03 \text{ g NH}_3}{1 \cancel{\text{mol NH}_3}} = 17.0 \text{ g NH}_3$$

 **Sense check:** 1.50 mol of H₂ makes 1.00 mol of NH₃ (2 per 3), and a mole of NH₃ is only 17 g, so 33.6 L of gas yields just 17.0 g. ✓

✓ Worked example 1: the route on the map



given: 33.6 L H₂ at STP · found: 17.0 g NH₃



🧠 **Sense check:** A is H₂, B is NH₃. Three arrows, one per method step: the molar volume, the mole ratio, the molar mass. ✓

Worked example 2



given: 130. g NaN_3 · wanted: L N_2 at STP

An airbag inflates when sodium azide decomposes to sodium metal and nitrogen gas. A **130.-g charge of NaN_3** fires. What volume of N_2 , at STP, does it release? (NaN_3 65.02 g/mol)

Write the route first: g $\text{NaN}_3 \rightarrow$ mol $\text{NaN}_3 \rightarrow$ mol $\text{N}_2 \rightarrow$ L N_2 .

✓ Worked example 2: solution



given: 130. g NaN₃ · wanted: L N₂ at STP

Step 1 · To moles

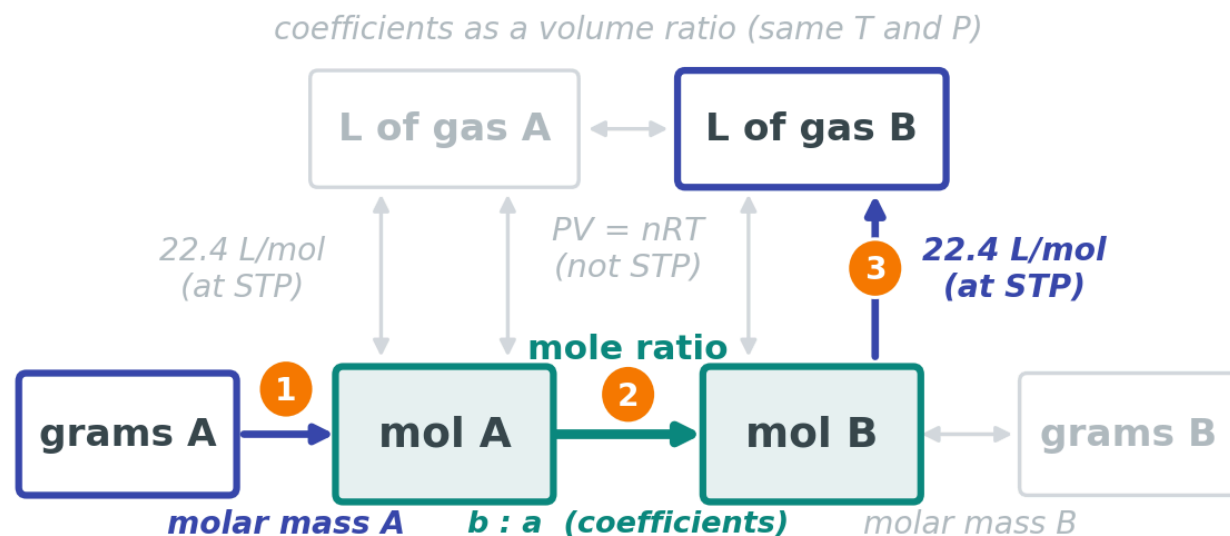
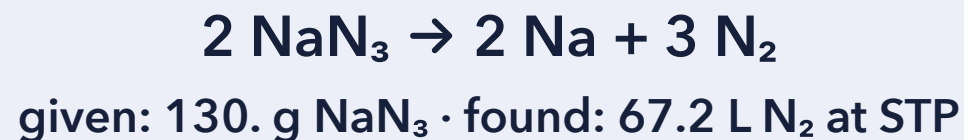
Step 2 · Moles → moles

Step 3 · To the wanted unit

$$130. \text{ g NaN}_3 \times \frac{1 \text{ mol NaN}_3}{65.02 \text{ g NaN}_3} \times \frac{3 \text{ mol N}_2}{2 \text{ mol NaN}_3} \times \frac{22.4 \text{ L N}_2}{1 \text{ mol N}_2} = 67.2 \text{ L N}_2$$

 **Sense check:** Two moles of solid NaN₃ (130 g) become three moles of N₂ gas: 67 L, enough to fill the bag. A small mass makes a large volume. ✓

✓ Worked example 2: the route on the map



🧠 **Sense check:** A is NaN_3 , B is N_2 . Grams enter through the molar mass; liters leave through the molar volume, because the gas is at STP. ✓

Your turn: hydrogen peroxide



given: 17.0 g H_2O_2 · wanted: L O_2 at STP (H_2O_2 34.02 g/mol)

Hydrogen peroxide decomposes to water and oxygen gas. Starting from **17.0 g H_2O_2** :

$$17.0 \text{ g } \text{H}_2\text{O}_2 \times \frac{1 \text{ mol } \text{H}_2\text{O}_2}{34.02 \text{ g } \text{H}_2\text{O}_2} \times \frac{\text{ } \text{mol } \text{O}_2}{\text{ } \text{mol } \text{H}_2\text{O}_2} \times \frac{\text{ } \text{L } \text{O}_2}{1 \text{ mol } \text{O}_2} = \text{ } \text{L } \text{O}_2$$

Fill the mole ratio from the coefficients and the molar-volume factor, then compute.

$$17.0 \text{ g } \text{H}_2\text{O}_2 \times \frac{1 \text{ mol } \text{H}_2\text{O}_2}{34.02 \text{ g } \text{H}_2\text{O}_2} \times \frac{1 \text{ mol } \text{O}_2}{2 \text{ mol } \text{H}_2\text{O}_2} \times \frac{22.4 \text{ L } \text{O}_2}{1 \text{ mol } \text{O}_2} = \mathbf{5.60 \text{ L } \text{O}_2}$$

Where this goes wrong

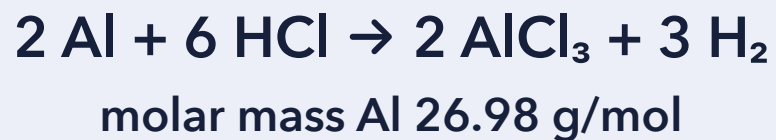


given: 17.0 g H_2O_2 · correct: 17.0 g $\rightarrow$ 0.500 mol $\rightarrow$ 0.250 mol $\text{O}_2 \rightarrow$ 5.60 L

Skipping the mole ratio. $17.0/34.02 \times 22.4 = 11.2$ L assumes one O_2 per H_2O_2 . The equation gives 2 H_2O_2 : 1 O_2 , and only the mole ratio switches substances.

Grams into the mole ratio. $17.0 \times \frac{1}{2} \times 22.4 = 190$ L skips the molar mass. Coefficients count moles, not grams: convert first, $17.0 \text{ g} \div 34.02 \text{ g/mol} = 0.500 \text{ mol}$.

Stopping at moles. $17.0/34.02 \div 2 = 0.250$ is moles of O_2 , not liters. Finish with the molar volume: $0.250 \text{ mol} \times 22.4 \text{ L/mol} = 5.60 \text{ L}$.

 Practice 1

Aluminum reacts with hydrochloric acid, releasing hydrogen gas. **8.10 g of Al** reacts completely. What volume of H_2 , in L, forms at STP?

- A. 6.72
- B. 0.450
- C. 4.48
- D. 10.1

✓ Practice 1 answer: D



given: 8.10 g Al · wanted: L H₂ at STP

$$8.10 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol H}_2}{2 \text{ mol Al}} \times \frac{22.4 \text{ L H}_2}{1 \text{ mol H}_2} = 10.1 \text{ L H}_2 \rightarrow \text{answer D}$$

A skipped the mole ratio: $8.10/26.98 \times 22.4 = 6.72$. B stopped at moles: $8.10/26.98 \times 3/2 = 0.450 \text{ mol}$, not liters. C flipped the mole ratio: $8.10/26.98 \times 2/3 \times 22.4 = 4.48$, as if fewer moles of H₂ formed than Al reacted.

 **Sense check:** 0.300 mol Al gives more moles of H₂ (3 per 2) = 0.450 mol, and each mole of gas is 22.4 L: about 10 L. 10.1 L ✓

Worked example 3



given: 42.0 g NaHCO_3 · wanted: L CO_2 at 27 °C, 1.00 atm

Baking soda decomposes on heating, releasing CO_2 . **42.0 g of NaHCO_3** decomposes, and the CO_2 is collected at **27 °C and 1.00 atm**, not STP. What volume forms? (NaHCO_3 84.01 g/mol)

The conditions are not STP, so 22.4 L/mol does not apply.

✓ Worked example 3: solution



given: 42.0 g NaHCO₃ · wanted: L CO₂ at 27 °C (300.15 K), 1.00 atm

Step 1 · To moles

Step 2 · Moles → moles

$$42.0 \text{ g NaHCO}_3 \times \frac{1 \text{ mol NaHCO}_3}{84.01 \text{ g NaHCO}_3} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol NaHCO}_3} = 0.250 \text{ mol CO}_2$$

The 22.4 shortcut

$$0.250 \text{ mol CO}_2 \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 5.60 \text{ L} \times$$

 **Sense check:** At STP the 0.250 mol would occupy 5.60 L. The gas is warmer than STP, so its true volume must be larger. ✓

✓ Worked example 3: solving $PV = nRT$



0.250 mol CO_2 · collected at 27 °C (300.15 K), 1.00 atm, not STP

Step 3 · To the wanted unit

Solve $PV = nRT$ for volume, with T in kelvin (27 + 273.15 = 300.15 K):

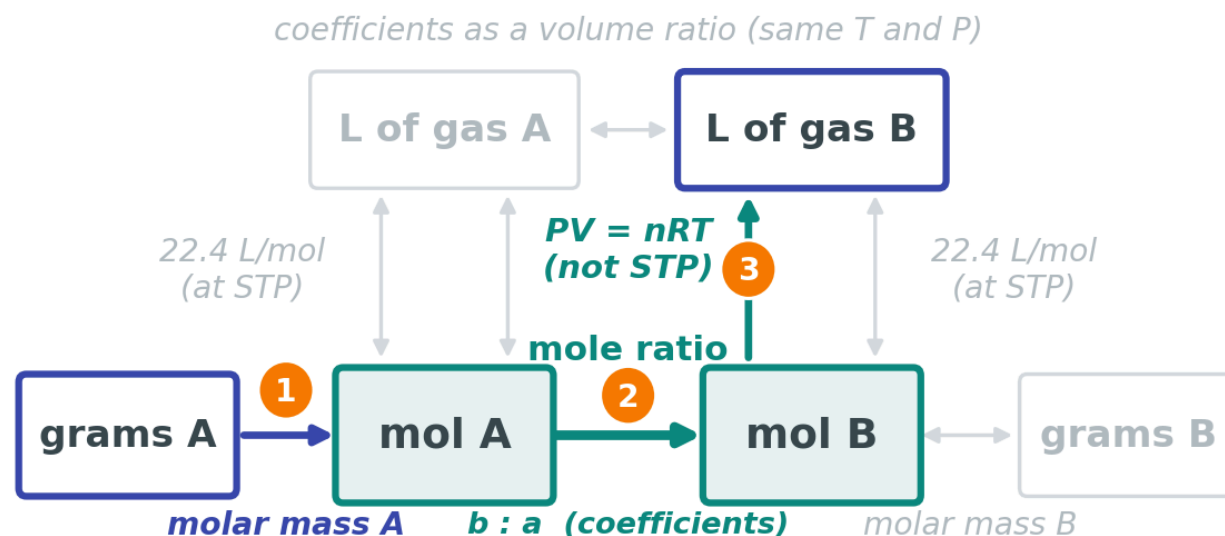
$$V = \frac{nRT}{P} = \frac{(0.250)(0.08206)(300.15)}{1.00} = \mathbf{6.16 \text{ L CO}_2}$$

🧠 **Sense check:** Warmer than STP, so each mole spreads past 22.4 L: 6.16 L exceeds the 5.60 L the shortcut gave. ✓

✓ Worked example 3: the route on the map



given: 42.0 g NaHCO_3 · found: 6.16 L CO_2 at 27 °C, 1.00 atm



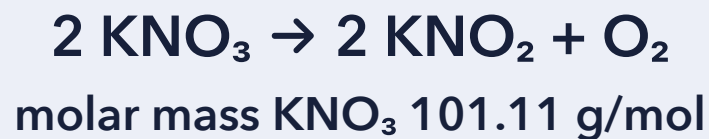
🧠 **Sense check:** Arrows 1 and 2 match an STP problem. Off STP, arrow 3 is $PV = nRT$, not 22.4 L/mol. ✓

Take-home: 22.4 L/mol is an STP-only shortcut

at STP: 22.4 L/mol · off STP: $V = nRT/P$

42.0 g $\text{NaHCO}_3 \rightarrow 0.250 \text{ mol CO}_2 \cdot 5.60 \text{ L (shortcut X) vs } 6.16 \text{ L (PV = nRT } \checkmark)$

The molar volume is $PV = nRT$ frozen at 0 °C and 1 atm. Change either, and the volume per mole shifts. Off STP, reach for $PV = nRT$.

 Practice 2

40.0 g of KNO_3 decomposes, and the O_2 is collected at **127 °C and 1.00 atm**. What volume of O_2 , in L, forms?

- A. 6.50
- B. 4.43
- C. 0.198
- D. 2.06

✓ Practice 2 answer: A



given: 40.0 g KNO_3 · wanted: L O_2 at 127 °C (400.15 K), 1.00 atm

$$40.0 \text{ g } \cancel{\text{KNO}_3} \times \frac{1 \cancel{\text{ mol KNO}_3}}{101.11 \text{ g } \cancel{\text{KNO}_3}} \times \frac{1 \text{ mol O}_2}{2 \cancel{\text{ mol KNO}_3}} = 0.198 \text{ mol O}_2$$

$$V = \frac{nRT}{P} = \frac{(0.198)(0.08206)(400.15)}{1.00} = 6.50 \text{ L O}_2 \rightarrow \text{answer A}$$

B used 22.4 L/mol off STP: $0.198 \times 22.4 = 4.43$. C stopped at moles: 0.198 mol O_2 is the amount, not the volume it fills. D forgot to convert to kelvin: $(0.198)(0.08206)(127) = 2.06$.

 **Sense check:** Warmer than STP, so each mole exceeds 22.4 L; but only half the KNO_3 becomes O_2 (2 : 1), landing at 6.50 L. ✓

Check yourself

1. A gas volume is given at STP. Name the single factor that turns it into moles, and say why 22.4 L/mol works for any gas.
2. The same reaction runs at 100 °C instead of STP. Which step of the route changes, and what replaces the molar volume?

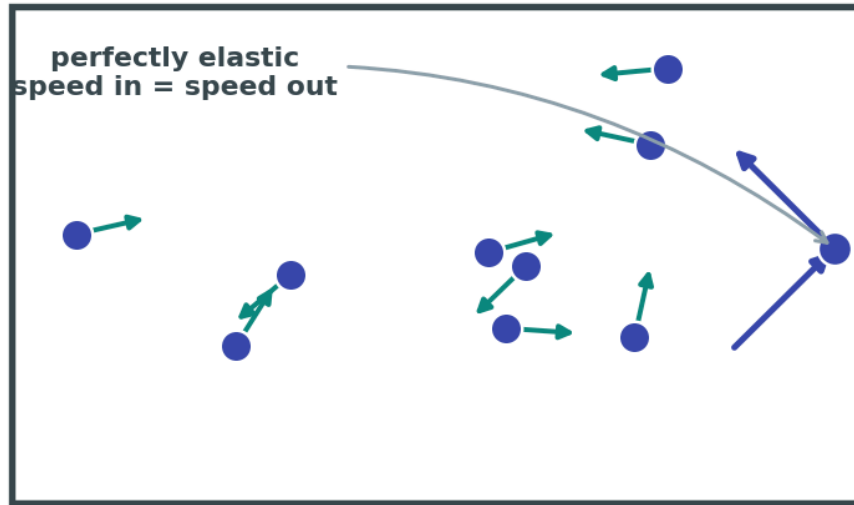
Cool a gas enough and it condenses to a liquid, then freezes to a solid: the same moles packed into a fixed volume. The forces that hold those particles together are the **intermolecular forces** of liquids and solids.

8 · Kinetic-Molecular Theory

State the five postulates of kinetic-molecular theory, use them to explain gas pressure and the gas laws, and compare gases at one temperature: same average kinetic energy, and the lighter gas moves faster.

Inside a balloon

A gas: tiny particles in ceaseless random motion



Pressure = countless collisions with the walls

Kinetic-molecular theory

- 1 Many tiny particles in constant, random, straight-line motion
- 2 Their volume is negligible: the container is mostly empty space
- 3 Collisions are perfectly elastic: no net kinetic energy lost
- 4 No forces between particles except during a collision
- 5 Average kinetic energy is proportional to Kelvin temperature

A balloon looks still. Inside, trillions of trillions of molecules fly in every direction and hammer the rubber without pause. That constant hammering is the pressure holding the balloon open.

Five postulates

1. Particles in constant, random, straight-line motion.
2. Their own volume is negligible.
3. Collisions are perfectly elastic.
4. No forces between particles.
5. Average kinetic energy is proportional to Kelvin temperature.

memory hook: Many Tiny Elastic Free Hot

many particles in motion · tiny volume · elastic collisions · free of forces · hot: KE follows Kelvin T

Where pressure comes from

pressure = total force of wall collisions ÷ wall area

more particles, less room, or faster particles $\Rightarrow$ more hits per second $\Rightarrow$ higher P

Each particle that strikes a wall gives it a tiny push. Pressure is the total of those pushes spread over the wall's area. More collisions, or harder ones, means more pressure.

One model, every gas law

pressure = hits per second $\times$ force per hit
every gas law changes one of the two

Each law is about wall hits.

- Boyle: smaller V, more frequent hits, higher P.
- Charles: higher T, harder hits, V grows.
- Avogadro: more particles, more hits, more V.
- Dalton: gases ignore each other; pressures add.

Temperature is average kinetic energy

$$KE_{\text{avg}} = (3/2)kT \text{ per molecule} = (3/2)RT \text{ per mole}$$

$k = 1.38 \times 10^{-23} \text{ J/K}$ · $R = 8.314 \text{ J/mol}\cdot\text{K}$ · T in kelvin · no mass term: the gas's identity never appears

Raising the temperature speeds the particles up. Temperature measures their average kinetic energy. At one temperature, a heavy gas and a light gas hold the same average kinetic energy.

Same energy, different speed

$$u_{\text{rms}} = \sqrt{(3RT / M)}$$

same T, same KE · smaller M, larger speed · speed ratio of two gases = $\sqrt{(M_{\text{heavy}} / M_{\text{light}})}$

Kinetic energy is $\frac{1}{2}mu^2$. Two gases with equal average kinetic energy but different masses cannot share a speed: the lighter one moves faster.

memory hook: same temperature, same average KE; lighter runs faster

helium outruns xenon at 300 K not because it carries more energy, but because it weighs less

Graham's law: lighter escapes faster

$$\text{rate}_A \div \text{rate}_B = \sqrt{M_B / M_A}$$

pinhole leak, same T · heavier gas on top under the root · H₂ vs CO₂: $\sqrt{44.01 \div 2.016} = 4.67$

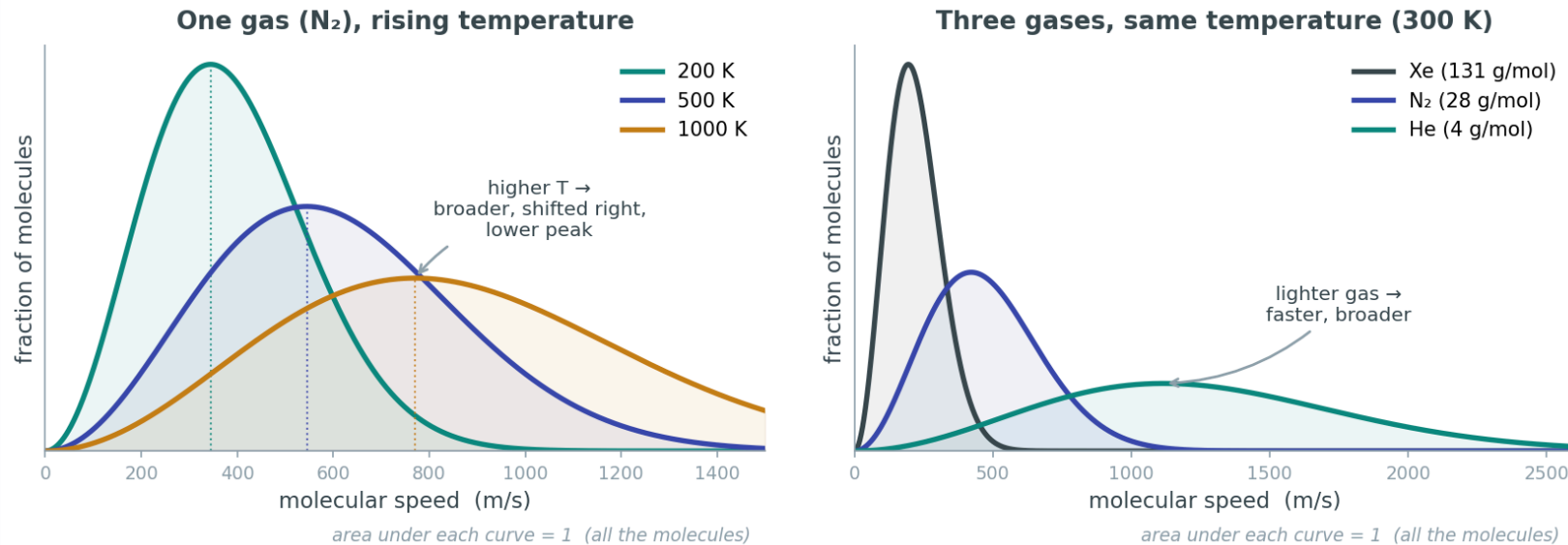
Effusion is a gas leaking through a pinhole. Faster molecules reach it more often, so the rate falls with the square root of molar mass. Hydrogen leaks nearly five times faster than carbon dioxide.

memory hook: light leaks first

smaller M, faster leak · a ratio below 1 for the lighter gas means the fraction is upside down

The Maxwell-Boltzmann distribution

The Maxwell-Boltzmann speed distribution



Molecules spread across a range of speeds. Heating shifts the curve right and flattens it; a heavier gas peaks at lower speed, taller and narrower. Every curve encloses the same area: all the molecules.



The method: compare gases with KMT

1. **Convert T to kelvin.**
2. **Compare energy.** Same T, same average KE, whatever the gas. KE follows T.
3. **Rank speeds.** Same T, smaller M is faster. Speed follows $\sqrt{T}$. The lighter gas also effuses faster.



Worked example 1: average kinetic energy

$$KE_{\text{avg}} = (3/2)RT \text{ per mole}$$

given: neon at 27 °C · R = 8.314 J/mol·K · wanted: KE_{avg} , then KE_{avg} after the Kelvin temperature doubles

A flask of neon sits at **27 °C**. Find the average kinetic energy per mole of its atoms. The gas is then heated until its **Kelvin temperature doubles**. Find the new value, and state how the average speed changes.

✓ Worked example 1: solution

$$KE_{\text{avg}} = (3/2)RT \text{ per mole}$$

neon · 27 °C · R = 8.314 J/mol·K

Step 1 · Convert T to kelvin

Step 2 · Compare energy

$$KE_{\text{avg}} = (3/2) \times 8.314 \text{ J/mol}\cdot\text{K} \times 300.15 \text{ K} = \mathbf{3743 \text{ J/mol}}$$

$$600.30 \text{ K: } KE_{\text{avg}} = (3/2) \times 8.314 \text{ J/mol}\cdot\text{K} \times 600.30 \text{ K} = \mathbf{7486 \text{ J/mol}}$$

Step 3 · Rank speeds: speed follows $\sqrt{T}$

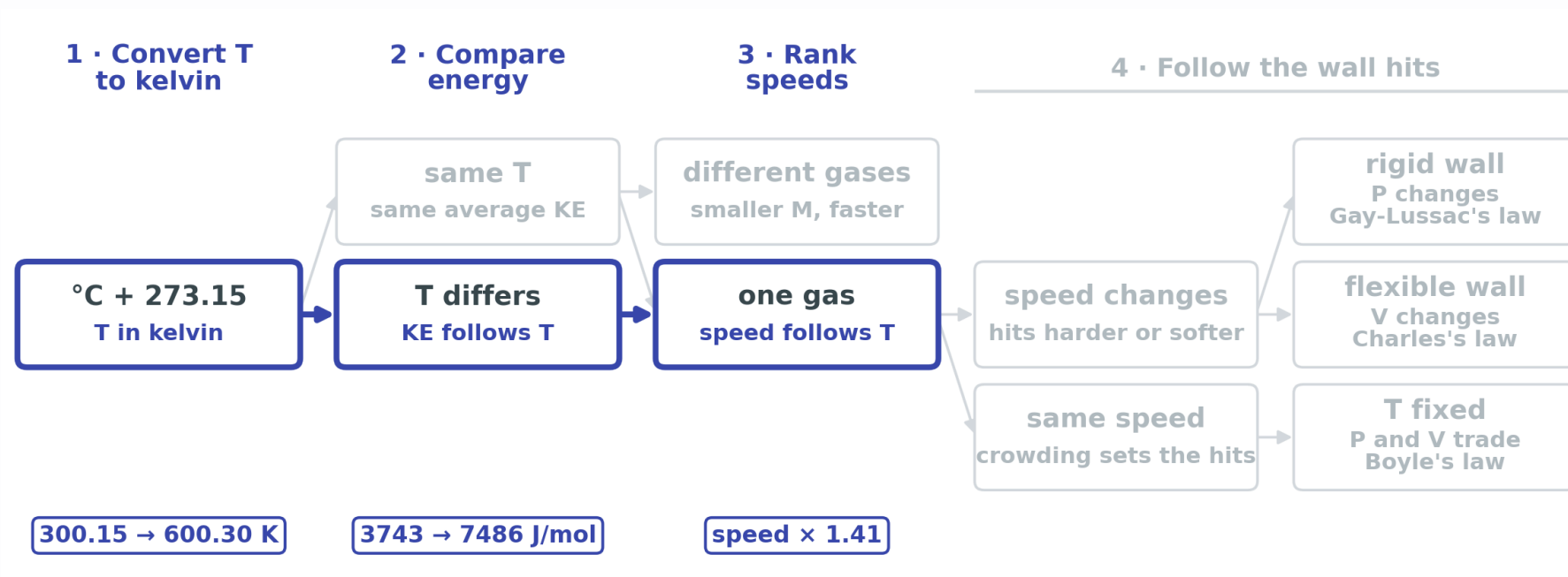
$$\text{energy: } 7486 \div 3743 = \mathbf{2.0} \cdot \text{speed: } \sqrt{2} = \mathbf{1.41}$$

 **Sense check:** Doubling in Celsius is not doubling: 27 → 54 °C is only 300.15 → 327.15 K, a factor of $327.15 \div 300.15 = 1.09$.

✓ Worked example 1: the route on the strip

$$KE_{avg} = (3/2)RT \text{ per mole}$$

given: neon at 27 °C · Kelvin T doubles · found: 3743 → 7486 J/mol · speed × 1.41



🧠 **Sense check:** One gas, and only T changed: the energy doubles with T, the speed grows by only $\sqrt{2}$. ✓



Worked example 2: ranking speeds

$$u_{\text{rms}} = \sqrt{(3RT / M)}$$

given: He, N₂, CO₂ · each at 25 °C · wanted: which gas has the most energy, and the order of average speeds

Three flasks at **25 °C** hold helium, nitrogen, and carbon dioxide. Which gas has the greatest average kinetic energy? Rank the three by average molecular speed.

A common first answer puts CO₂ fastest: the heaviest molecule carries the most energy. Test it.

✓ Worked example 2: solution

$$u_{\text{rms}} = \sqrt{(3RT / M)}$$

He 4.003 · N₂ 28.02 · CO₂ 44.01 g/mol · all at 25 °C

Step 1 · Convert T to kelvin

Step 2 · Compare energy

Step 3 · Rank speeds

M: He 4.003 < N₂ 28.02 < CO₂ 44.01 g/mol ⇒ speed: He > N₂ > CO₂

smaller M, larger speed · the heaviest gas is the slowest

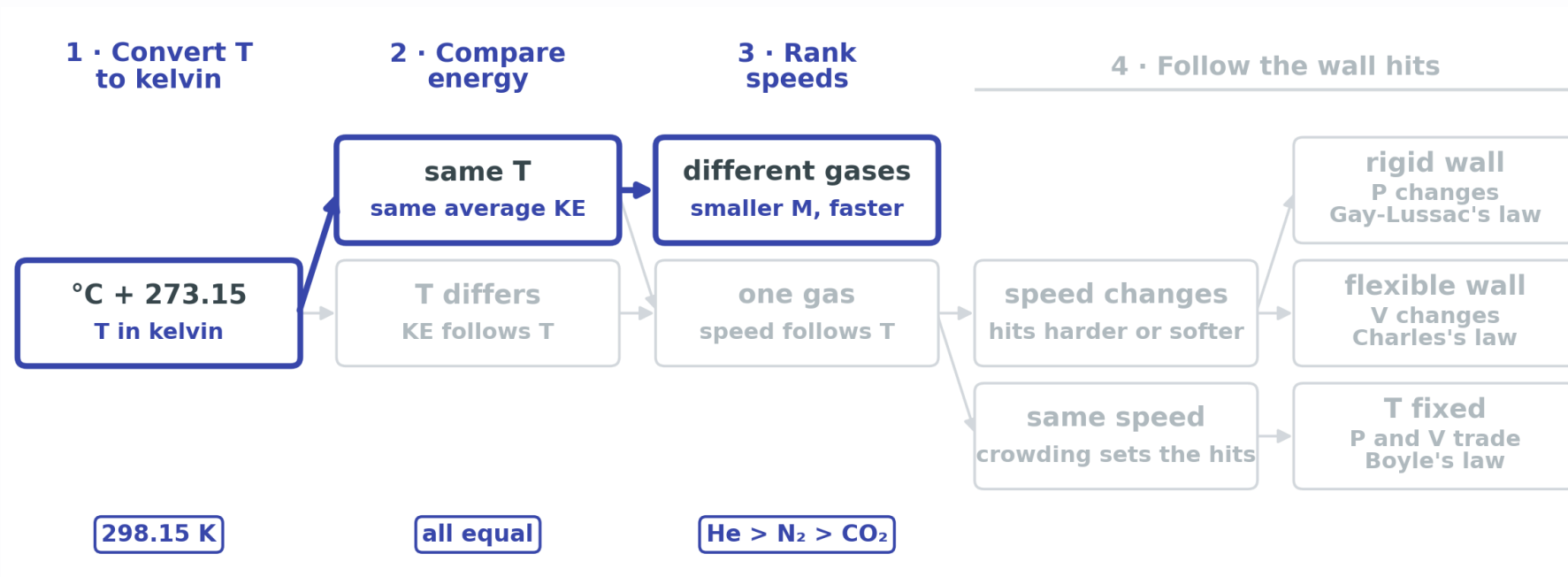
$$u(\text{He}) \div u(\text{CO}_2) = \sqrt{(44.01 \div 4.003)} = 3.32$$

 **Sense check:** Helium atoms move about three times faster than CO₂ molecules, yet carry the same average energy: less mass, more speed. The ratio needs no unit conversion; the molar masses cancel.

✓ Worked example 2: the route on the strip

$$u_{\text{rms}} = \sqrt{(3RT / M)}$$

given: He, N₂, CO₂ · each at 25 °C · found: equal average KE · speed He > N₂ > CO₂



🧠 **Sense check:** One temperature for all three flasks, so energy decides nothing. Molar mass alone ranks the speeds.



Your turn: helium and argon at 25 °C

$$u_{\text{rms}} = \sqrt{(3RT / M)}$$

given: He 4.003 · Ar 39.95 g/mol · both at 298.15 K · wanted: which has more energy, which is faster, and by what factor

same T $\Rightarrow$ average KE of He ——— average KE of Ar

$$u(\text{He}) \div u(\text{Ar}) = \sqrt{(\text{———} \div 4.003)} = \text{———}$$

Fill in the energy comparison, then the speed ratio.

$$\text{same T} \Rightarrow \text{average KE of He} = \text{average KE of Ar} \cdot u(\text{He}) \div u(\text{Ar}) = \sqrt{(39.95 \div 4.003)} = \mathbf{3.16}$$

 **Sense check:** Argon is ten times heavier ($39.95 \div 4.003 = 9.98$), so helium moves $\sqrt{10} = 3.16$ times faster. Same energy, less mass, more speed.

⚠ Where this goes wrong

Heavier means more energy. At the same T all gases share the same average KE. A heavier molecule moves slower to carry it.

Faster means more energy. Speed alone is not KE. A light, fast helium atom and a heavy, slow xenon atom at one temperature have equal average KE.

Doubling in Celsius. $KE \propto$ Kelvin temperature. Going $27 \rightarrow 54\text{ }^{\circ}\text{C}$ is $300.15 \rightarrow 327.15\text{ K}$, a factor of 1.09, not 2. Convert to kelvin first.

Speed factor = energy factor. Double the KE and the speed rises by only $\sqrt{2} = 1.41$, since $KE \propto u^2$. The two do not scale together.



Practice 1

$$KE_{\text{avg}} = (3/2)kT \text{ per molecule}$$

given: He, O₂, SO₂ · one mole each · same T = 350 K

Three flasks hold one mole each of He, O₂, and SO₂ at **350 K**. Which sample has the greatest average kinetic energy per molecule?

- A. He: its atoms move the fastest
- B. They are all equal
- C. SO₂: its molecules are the most massive
- D. It depends on the pressure in each flask

✓ Practice 1 answer: B

$$KE_{\text{avg}} = (3/2)kT: \text{no mass term, no pressure term}$$

same $T \Rightarrow$ same average KE for every gas, whatever its molar mass $\rightarrow$ answer B

Average kinetic energy depends only on Kelvin temperature, so all three are equal. A confuses speed with energy: He is the fastest, but that is how it carries the same KE with less mass. C assumes mass sets the energy; mass sets the speed. D adds a variable that never appears: $(3/2)kT$ has no pressure term.

 **Sense check:** At one temperature the energies match and the speeds differ. Only a change in temperature changes the average KE, and then for all three together.



Practice 2

$$KE_{\text{avg}} = (3/2)RT \text{ per mole} \cdot T \text{ in kelvin}$$

given: He at 10 °C · N₂ at 300 K · CH₄ at 30 °C · Xe at 60 °C · wanted: average KE, lowest to highest

Four flasks hold helium at **10 °C**, nitrogen at **300 K**, methane at **30 °C**, and xenon at **60 °C**. Rank the gases from lowest to highest average kinetic energy.

- A. All four are equal: average KE does not depend on the gas
- B. He < CH₄ < Xe < N₂
- C. Xe < N₂ < CH₄ < He
- D. He < N₂ < CH₄ < Xe

✓ Practice 2 answer: D

$$\text{He } 10 + 273.15 = 283.15 \text{ K} \cdot \text{N}_2 \text{ } 300 \text{ K} \cdot \text{CH}_4 \text{ } 30 + 273.15 = 303.15 \text{ K} \cdot \text{Xe } 60 + 273.15 = 333.15 \text{ K}$$

KE follows Kelvin T alone · lowest T, lowest KE → He < N₂ < CH₄ < Xe, answer D

A applied the same-temperature rule, but these flasks sit at four different temperatures. B compared the raw numbers, setting 300 K against Celsius readings; convert every T to kelvin first. C ranked by speed, lightest fastest, but speed is not energy: mass never enters $(3/2)RT$.

 **Sense check:** Xenon is the heaviest and slowest gas here, yet at 333.15 K it carries the most energy per molecule. Mass sets the speed; only T sets the average KE.



Practice 3

$$u_{\text{rms}} = \sqrt{(3RT / M)}$$

given: CH_4 , NH_3 , Ar , SO_2 · all at $100\text{ }^\circ\text{C}$ · wanted: the gas with the highest average speed

Four flasks at **$100\text{ }^\circ\text{C}$** hold methane, ammonia, argon, and sulfur dioxide. In which gas do the molecules move fastest on average?

- A. SO_2
- B. Ar
- C. CH_4
- D. NH_3
- E. All four are equal at one temperature

✓ Practice 3 answer: C

M: CH₄ 16.04 < NH₃ 17.03 < Ar 39.95 < SO₂ 64.07 g/mol

T = 100 + 273.15 = 373.15 K for all four · smallest M, largest speed → CH₄, answer C

$$u(\text{CH}_4) \div u(\text{NH}_3) = \sqrt{(17.03 \div 16.04)} = \mathbf{1.03} \cdot u(\text{CH}_4) \div u(\text{SO}_2) = \sqrt{(64.07 \div 16.04)} = \mathbf{2.00}$$

A picked the heaviest gas, which is the slowest. B picked the single atom, but Ar (39.95 g/mol) is heavier than the two light molecules. D is close, but 17.03 > 16.04, so ammonia trails methane by 3%. E is true of the average kinetic energy, not the speed.

 **Sense check:** Methane and ammonia are nearly tied; sulfur dioxide, four times heavier than methane, moves at half its speed.

 Practice 4

$$\text{rate}_A \div \text{rate}_B = \sqrt{M_B / M_A}$$

given: Ne and Kr · identical pinholes · same T · wanted: $\text{rate}(\text{Ne}) \div \text{rate}(\text{Kr})$

Neon and krypton leak through identical pinholes at **25 °C**. How many times faster does neon effuse than krypton?

- A. 4.15
- B. 2.04
- C. 0.491
- D. 0.241

✓ Practice 4 answer: B

$$\text{rate}(\text{Ne}) \div \text{rate}(\text{Kr}) = \sqrt{(M_{\text{Kr}} / M_{\text{Ne}})}$$

Ne 20.18 · Kr 83.80 g/mol · the heavier gas goes on top under the root

$$\text{rate}(\text{Ne}) \div \text{rate}(\text{Kr}) = \sqrt{(83.80 \div 20.18)} = \sqrt{4.15} = \mathbf{2.04} \rightarrow \text{answer B}$$

A took the molar-mass ratio without the root: $83.80 \div 20.18 = 4.15$. C inverted the fraction: $\sqrt{(20.18 \div 83.80)} = 0.491$, which would make the lighter gas slower. D inverted it and skipped the root: $20.18 \div 83.80 = 0.241$.

 **Sense check:** Krypton is four times heavier, yet neon leaks only twice as fast. The square root halves the effect, and the lighter gas always comes out above 1.

 Practice 5

$$u_{\text{rms}} = \sqrt{(3RT / M)}$$

given: He at 30 K · Ne at 200 K · Ar at 250 K · wanted: rms speed, slowest first

Helium at **30 K**, neon at **200 K**, and argon at **250 K** sit in three sealed flasks. Order them by rms speed, slowest first.

- A. Ar < He < Ne
- B. He < Ne < Ar
- C. Ar < Ne < He
- D. Ne < He < Ar

✓ Practice 5 answer: A

$$u_{\text{rms}} = \sqrt{(3RT / M)}: \text{speed follows } T \div M$$

He 4.003 · Ne 20.18 · Ar 39.95 g/mol · T and M both differ, so neither rule alone decides

$$T \div M: \text{Ar } 250 \div 39.95 = \mathbf{6.26} \cdot \text{He } 30 \div 4.003 = \mathbf{7.49} \cdot \text{Ne } 200 \div 20.18 = \mathbf{9.91}$$

→ Ar < He < Ne, answer A

B ranked by temperature alone, as if all three had one mass. C ranked by mass alone, lightest fastest, as if all three shared one temperature. D put M over T, the fraction under the root upside down, so the order runs backward.

🧠 **Sense check:** Helium is the lightest, yet at 30 K it trails neon: neon is 5.04 times heavier but 6.67 times hotter. ✓

Check yourself

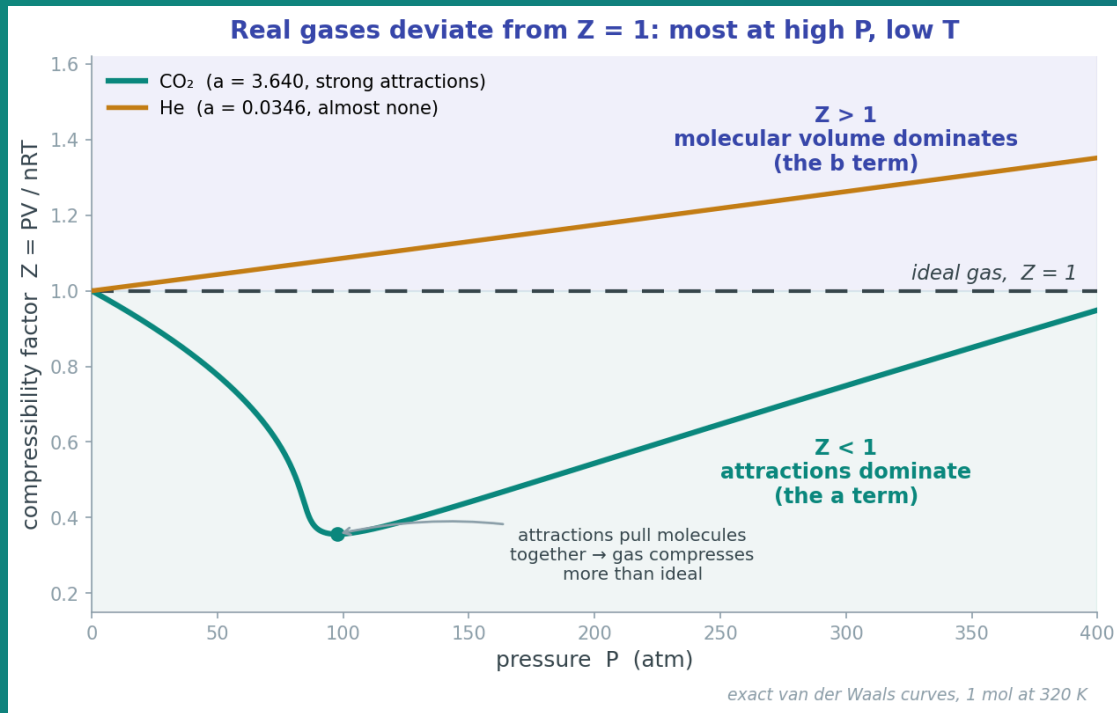
1. Two flasks at 25 °C hold helium and argon. Which gas has the greater average kinetic energy, and which has the greater average speed?
2. A gas is warmed so its Kelvin temperature doubles. By what factor does the average kinetic energy change, and by what factor does the average speed change?

The five postulates are the microscopic story behind every gas law. Give the particles real volume and real attractions, and this ideal model becomes the behavior of real gases.

9 · Real Gases and the van der Waals Equation

Explain why real gases deviate from $PV = nRT$ (and when), and use the van der Waals equation $(P + a \cdot n^2/V^2)(V - n \cdot b) = nRT$ to correct the pressure of a real gas for molecular volume (b) and intermolecular attractions (a).

When $PV = nRT$ fails



Air in a scuba tank sits at 200 atm. The molecules are so crowded that their size and mutual pull change the pressure. $PV = nRT$ is no longer exact.

The five postulates, revisited

$PV = nRT$ is built on all five kinetic-molecular postulates

**motion · elastic collisions · KE follows Kelvin T: hold for a real gas · tiny volume · free of forces:
approximations**

Every gas law so far, from Boyle to gas stoichiometry, rests on these five postulates. Two are approximations. Real molecules have size, and real molecules attract. When either matters, $P = nRT/V$ needs a correction.

Two false assumptions

$PV = nRT$ assumes point molecules that never attract
both hold well at everyday conditions · both fail when the gas is crowded or cold

ideal assumption	a real gas
zero molecular volume	has size; crowded at high P
no attractions	pulls together; slow at low T

Deviations grow at high P and low T

deviation grows when molecules are crowded or slow

**high P: crowded, so size and pull both count · low T: slow, so attractions take hold · low P and high T:
nearly ideal**

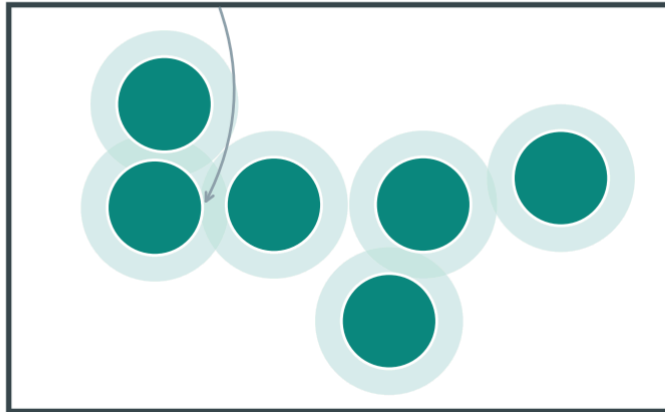
High pressure packs molecules close, so their volume and mutual pull matter. Low temperature slows them, so attractions take hold. At low P and high T the gas is nearly ideal.

Two corrections, one for each assumption

Two reasons real gases break $PV = nRT$

Molecules take up space: subtract $n \cdot b$

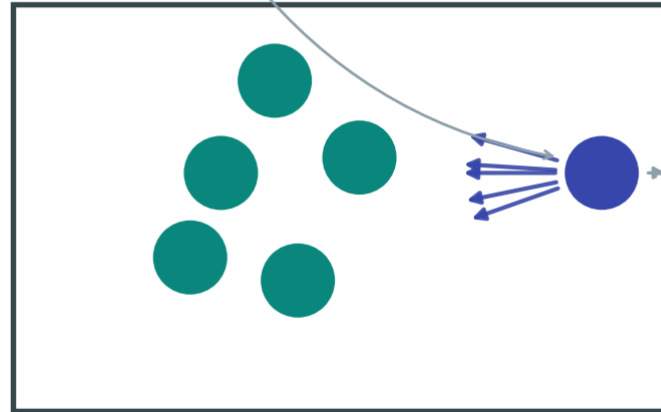
each molecule owns
a finite volume $\rightarrow b$



real free volume = $V - n \cdot b$ (not the full V)

Neighbors pull a molecule back: add $a \cdot n^2/V^2$

pulled inward by
neighbors $\rightarrow$ hits the
wall more softly



measured P is LOWER than ideal $\rightarrow$ add $a \cdot n^2/V^2$

Van der Waals kept $PV = nRT$ and patched each false assumption with one constant.

memory hook: a for attraction, b for bulk

a adds back the pressure attractions steal $\cdot$ b subtracts the space molecules fill

The van der Waals equation

$$(P + a \cdot n^2/V^2)(V - n \cdot b) = nRT$$

$a \cdot n^2/V^2$: pressure lost to attractions, added back · $n \cdot b$: volume the molecules fill, subtracted · a in $\text{L}^2 \cdot \text{atm}/\text{mol}^2$ · b in L/mol

Larger molecules have a larger b . Stronger intermolecular attractions give a larger a . Set both to zero and the equation is $PV = nRT$ again.

a : He 0.0346 · N₂ 1.352 · CO₂ 3.640 · H₂O 5.536 $\text{L}^2 \cdot \text{atm}/\text{mol}^2$

weak attractions to strong · b : He 0.0238 · N₂ 0.0387 · CO₂ 0.0427 L/mol

The compressibility factor Z

$$Z = PV / nRT$$

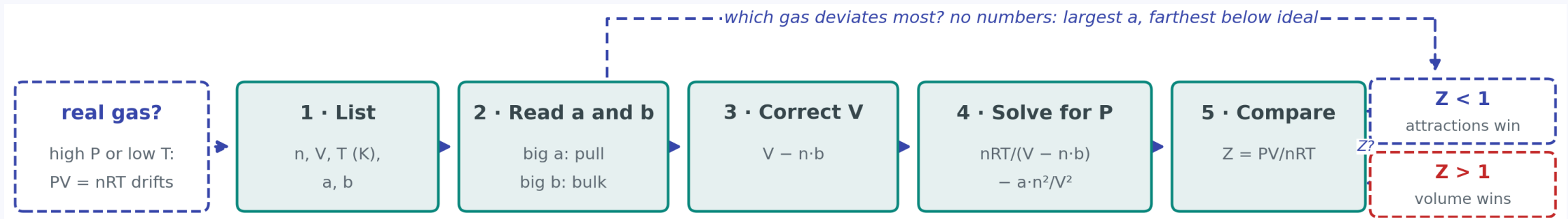
ideal gas: $Z = 1$ · $Z < 1$: attractions dominate, PV falls short · $Z > 1$: molecular volume dominates, PV overshoots

At low-to-moderate pressure attractions win and Z sits below 1. At very high pressure the molecules' own volume wins and Z climbs above 1.



The method

1. **List** n , V , T (kelvin), a , b .
2. **Read a and b** . Big a , strong attraction; big b , bulk.
3. **Correct the volume:** $V - n \cdot b$.
4. **Solve for P :** $nRT/(V - n \cdot b) - a \cdot n^2/V^2$.
5. **Compare with ideal:** $Z = PV/nRT$.



Worked example 1: which gas deviates most

a: He 0.0346 · N₂ 1.352 · CO₂ 3.640 L²·atm/mol²

given: 1 mol of each at the same moderate P and T · wanted: the gas farthest from $PV = nRT$

Three cylinders hold one mole each of helium, nitrogen, and carbon dioxide at the same moderate pressure and temperature. Decide which gas departs most from ideal behavior.

A common first answer picks helium, the smallest molecule, as the easiest to squeeze. Test it against the constants.

✓ Worked example 1: solution

Step 1 • List

n, P, T identical for all three • only a and b differ

a : He 0.0346 • N₂ 1.352 • CO₂ 3.640 L²·atm/mol² • b : He 0.0238 • N₂ 0.0387 • CO₂ 0.0427 L/mol

Step 2 • Read a and b

deviation: CO₂ > N₂ > He

largest a , largest pull, largest gap from $PV = nRT$ • helium, with almost no attractions, stays nearly ideal

 **Sense check:** The smallest molecule is not the one that deviates most. Helium is tiny and barely attracts; it is the closest to ideal of all gases.





Worked example 2: CO₂ under pressure

$$(P + a \cdot n^2/V^2)(V - n \cdot b) = nRT$$

given: 1.00 mol CO₂ · 0.500 L · 300. K · a = 3.640 L²·atm/mol² · b = 0.04267 L/mol · wanted: P

1.00 mol of CO₂ is squeezed into **0.500 L** at **300. K**. Find the pressure from the van der Waals equation and compare it with the ideal-gas value.

✓ Worked example 2: solution

Step 1 • List

$$n = 1.00 \text{ mol} \cdot V = 0.500 \text{ L} \cdot T = 300. \text{ K}$$

$$a = 3.640 \text{ L}^2 \cdot \text{atm/mol}^2 \cdot b = 0.04267 \text{ L/mol} \cdot R = 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K}$$

Step 2 • Read a and b

Step 3 • Correct the volume

Step 4 • Solve for P

$$P = \frac{1.00 \times 0.08206 \times 300.}{0.500 - 0.04267} - \frac{3.640 \times 1.00^2}{0.500^2} = 53.83 - 14.56 = \mathbf{39.3 \text{ atm}}$$

Step 5 • Compare with ideal

$$\text{ideal } P = 24.618 \div 0.500 = \mathbf{49.2 \text{ atm}} \cdot Z = 39.3 \times 0.500 \div 24.618 = \mathbf{0.798}$$

 **Sense check:** Z below 1: attractions dominate. The real gas exerts about 10 atm less than ideal.

✓ Worked example 2: the route on the map

$$P = nRT/(V - n \cdot b) - a \cdot n^2/V^2$$

given: 1.00 mol CO₂ · 0.500 L · 300. K · found: P = 39.3 atm · ideal 49.2 atm · Z = 0.798



🧠 **Sense check:** All five steps run in order. Step 5 reads the verdict: Z below 1, so attractions win.

Your turn: nitrogen under pressure

$$P = nRT/(V - n \cdot b) - a \cdot n^2/V^2$$

given: 2.00 mol N₂ · 1.00 L · 300. K · a = 1.352 L²·atm/mol² · b = 0.0387 L/mol · ideal P = 49.2 atm ·

wanted: is the real P above or below, and which term wins

$$V - n \cdot b = 1.00 \text{ L} - 2.00 \times 0.0387 \text{ L} = \underline{\hspace{1cm}} \text{ L} \cdot nRT \div (V - n \cdot b) = 49.236 \div \underline{\hspace{1cm}} = \underline{\hspace{1cm}} \text{ atm}$$

$$a \cdot n^2/V^2 = 1.352 \times 2.00^2 \div 1.00^2 = \underline{\hspace{1cm}} \text{ atm} \cdot P = \underline{\hspace{1cm}} - \underline{\hspace{1cm}} = \underline{\hspace{1cm}} \text{ atm}$$

$$V - n \cdot b = 1.00 - 0.0774 = \mathbf{0.9226 \text{ L}} \cdot 49.236 \div 0.9226 = \mathbf{53.37 \text{ atm}}, \text{ up } 4.13 \text{ from ideal}$$

$$a \cdot n^2/V^2 = 1.352 \times 4.00 = \mathbf{5.41 \text{ atm}} \cdot P = 53.37 - 5.41 = \mathbf{48.0 \text{ atm}}, \text{ below ideal}$$

 **Sense check:** The attraction term (5.41 atm down) beats the volume term (4.13 atm up), so the real P sits below ideal: $Z = 48.0 \times 1.00 \div 49.236 = 0.974$. Nitrogen's small a keeps it close.

⚠ Where this goes wrong

Dropping the $n \cdot b$ from V . Using nRT/V instead of $nRT/(V - n \cdot b)$ ignores the space molecules fill. The volume correction vanishes.

Adding the a term to P . Attractions make the measured pressure lower, so $a \cdot n^2/V^2$ is subtracted when solving for P . The plus sign lives inside the equation's first factor, not in the answer.

Assuming Z is always above 1. At low-to-moderate pressure attractions win and $Z < 1$. Only at very high pressure does molecular volume push Z above 1.

Expecting big deviations at low P and high T . There the gas is nearly ideal. Deviations are largest at high P and low T .



Practice 1

a ($\text{L}^2 \cdot \text{atm} / \text{mol}^2$): NH_3 4.225 · Kr 2.318 · CH_4 2.283 · H_2 0.2476

b (L / mol): NH_3 0.03707 · Kr 0.03978 · CH_4 0.04278 · H_2 0.02661 · wanted: the largest drop below nRT/V

Identical tanks hold equal amounts of four gases at room temperature and moderate pressure. Which gas's measured pressure falls farthest below the ideal-gas value?

- A. Kr
- B. NH_3
- C. CH_4
- D. H_2

✓ Practice 1 answer: B

moderate P: attractions control the deviation, so compare a

a: NH_3 4.225 > Kr 2.318 > CH_4 2.283 > H_2 0.2476 $\text{L}^2\cdot\text{atm}/\text{mol}^2$ · largest a, strongest pull, lowest real P → answer B

A picked the heaviest atom: molar mass is not one of the two corrections. C read the b column: CH_4 has the largest b, 0.04278 L/mol, and the b term pushes P up, not down. D picked the smallest a: H_2 is the most nearly ideal of the four.

 **Sense check:** Ammonia molecules hydrogen-bond to one another, which gives NH_3 the largest a here. Strong attractions, large a, Z farthest below 1.





Practice 2

$$(P + a \cdot n^2/V^2)(V - n \cdot b) = nRT$$

given: 1.00 mol CH₄ · 0.500 L · 350 K · a = 2.283 L²·atm/mol² · b = 0.04278 L/mol · wanted: P

A steel cylinder holds **1.00 mol** of methane in **0.500 L** at **350 K**. The ideal gas law predicts **57.4 atm**. Is the real pressure above or below that, and which correction wins?

- A. Above ideal, 62.8 atm: only the volume term counts
- B. Below ideal, 48.3 atm: only the attraction term counts
- C. Below ideal, 53.7 atm: the attraction term outweighs the volume term
- D. Above ideal, 71.9 atm: the attraction term is added to P

✓ Practice 2 answer: C

$$V - n \cdot b = 0.500 - 1.00 \times 0.04278 = 0.4572 \text{ L}$$

$$nRT = 1.00 \times 0.08206 \times 350 = 28.72 \text{ L} \cdot \text{atm} \cdot a \cdot n^2/V^2 = 2.283 \times 1.00^2 \div 0.500^2 = 9.132 \text{ atm}$$

$$P = \frac{28.72}{0.4572} - 9.132 = 62.82 - 9.13 = \mathbf{53.7 \text{ atm}}, \text{ below } 57.4 \rightarrow \text{answer C}$$

The volume term lifts P by $62.82 - 57.44 = 5.38 \text{ atm}$; the attraction term drops it by 9.13 atm , so the net is 3.76 atm below ideal. A kept only the volume correction: 62.8 atm . B kept only the attraction term: $57.44 - 9.13 = 48.3 \text{ atm}$. D added the a term instead of subtracting it: $62.82 + 9.13 = 71.9 \text{ atm}$.

 **Sense check:** Real P below ideal, as expected at moderate pressure: $Z = 53.7 \times 0.500 \div 28.72 = 0.93$, attractions dominate.





Guided example: helium at high pressure

$$(P + a \cdot n^2/V^2)(V - n \cdot b) = nRT$$

given: 2.00 mol He · 0.250 L · 300. K · $a = 0.0346 \text{ L}^2 \cdot \text{atm}/\text{mol}^2$ · $b = 0.0238 \text{ L}/\text{mol}$ · wanted: P, and which correction wins

A **0.250 L** steel cartridge holds **2.00 mol** of helium at **300. K**, as crowded as the gas in a full scuba tank. Find P from the van der Waals equation and compare it with the ideal value.

✓ Guided example: solution

Step 1 • List

$$n = 2.00 \text{ mol} \cdot V = 0.250 \text{ L} \cdot T = 300. \text{ K} \cdot R = 0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$$

$$a = 0.0346 \text{ L}^2\cdot\text{atm}/\text{mol}^2 \cdot b = 0.0238 \text{ L}/\text{mol} \cdot nRT = 2.00 \times 0.08206 \times 300. = 49.236 \text{ L}\cdot\text{atm}$$

Step 2 • Read a and b

Step 3 • Correct the volume

Step 4 • Solve for P

$$P = 49.236 \div (0.250 - 0.0476) - (0.0346 \times 2.00^2) \div 0.250^2 = 243.26 - 2.21 = \mathbf{241 \text{ atm}}$$

Step 5 • Compare with ideal

$$\text{ideal } P = nRT \div V = 49.236 \div 0.250 = \mathbf{197 \text{ atm}} \cdot Z = 241.05 \times 0.250 \div 49.236 = \mathbf{1.22}$$

🧠 **Sense check:** Z above 1: volume wins. The b term adds 46.3 atm; the a term takes back only 2.21 atm.





Practice 3

$$(P + a \cdot n^2/V^2)(V - n \cdot b) = nRT$$

given: 263 g Xe · 0.800 L · 315 K · $a = 4.192 \text{ L}^2 \cdot \text{atm/mol}^2$ · $b = 0.05156 \text{ L/mol}$ · wanted: P

Satellite ion thrusters run on xenon stored as a dense gas. A **0.800 L** propellant tank is loaded with **263 g** of xenon at **315 K**. Find its pressure, in atm.

- A. 61.2
- B. 101
- C. 53.3
- D. 64.7
- E. 48.0

✓ Practice 3 answer: E

$$n = 263 \text{ g} \div 131.29 \text{ g/mol} = 2.003 \text{ mol} \cdot V - n \cdot b = 0.800 - 0.1033 = 0.6967 \text{ L}$$

$$n \cdot b = 2.003 \times 0.05156 = 0.1033 \text{ L} \cdot nRT = 2.003 \times 0.08206 \times 315 = 51.78 \text{ L} \cdot \text{atm} \cdot a \cdot n^2/V^2 = 4.192 \times$$

$$2.003^2 \div 0.800^2 = 26.28 \text{ atm}$$

$$P = \frac{51.78}{0.6967} - 26.28 = 74.32 - 26.28 = \mathbf{48.0} \text{ atm} \rightarrow \text{answer E}$$

A left n unsquared: $74.32 - 13.12 = 61.2$. B added the a term: $74.32 + 26.28 = 101$. C divided by V , not V^2 : $74.32 - 21.02 = 53.3$. D skipped both corrections: $51.78 \div 0.800 = 64.7$, the ideal pressure.

🧠 **Sense check:** $Z = 48.0 \times 0.800 \div 51.78 = 0.742$. Xenon's large a pulls the real pressure 16.7 atm below ideal.



Check yourself

1. In $(P + a \cdot n^2/V^2)(V - n \cdot b) = nRT$, state what a and b each correct for, and whether a larger value means bigger molecules or stronger attractions.
2. A gas at moderate pressure has $Z = PV/nRT = 0.93$. Which effect dominates, and is its measured pressure above or below the ideal value?

Real gases obey $PV = nRT$ best when dilute and warm. The same attractions that pull Z below 1 are the intermolecular forces that, taken further, condense a gas into a liquid.

Can you...?

- ☐ state the conditions of STP and use the molar volume (22.4 L/mol) to convert between moles and volume of a gas?
- ☐ apply the combined gas law to relate the pressure, volume, and temperature of a gas sample?
- ☐ use the ideal gas law $PV = nRT$ to solve for any variable, a gas density, or a molar mass?
- ☐ apply Dalton's law of partial pressures, including a gas collected over water?
- ☐ carry mole ratios through gas-stoichiometry problems at STP and non-STP conditions?
- ☐ use Graham's law to compare effusion rates of gases from their molar masses?

If any box stays empty, the practice site has a drill for it. 🩹