

Gases

Preparation for General Chemistry · Dr. Karmach

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

- Describe gas behavior with kinetic-molecular theory and convert pressure among atm, mmHg, kPa, and psi
- Apply Boyle's, Charles's, Gay-Lussac's, Avogadro's, and the combined gas law to a fixed gas sample, with temperature in kelvin
- Use $PV = nRT$ to find pressure, volume, amount, or temperature, and the density or molar mass of a gas
- Convert between moles and liters of a gas at STP with 22.4 L/mol and carry gas volumes through stoichiometry
- Apply Dalton's law to gas mixtures, including a gas collected over water
- Compare gases at one temperature: same average kinetic energy, and the lighter gas moves, diffuses, and effuses faster

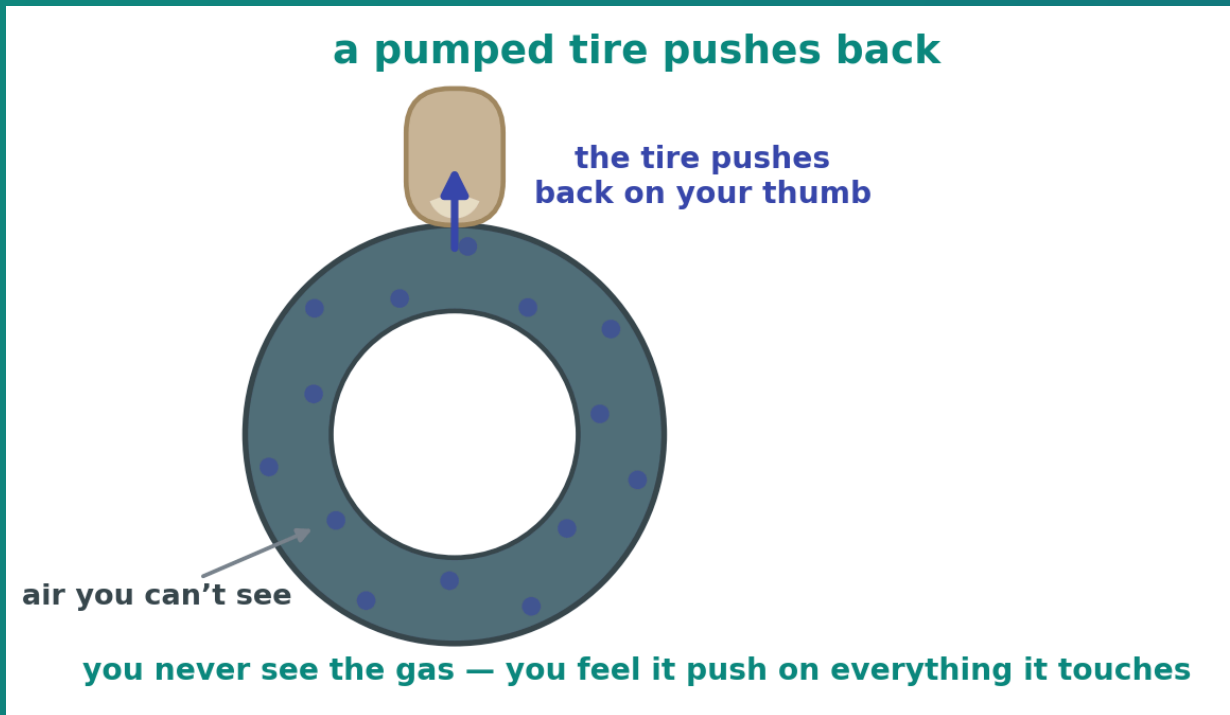
Today's route

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**

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, Pa, and psi.

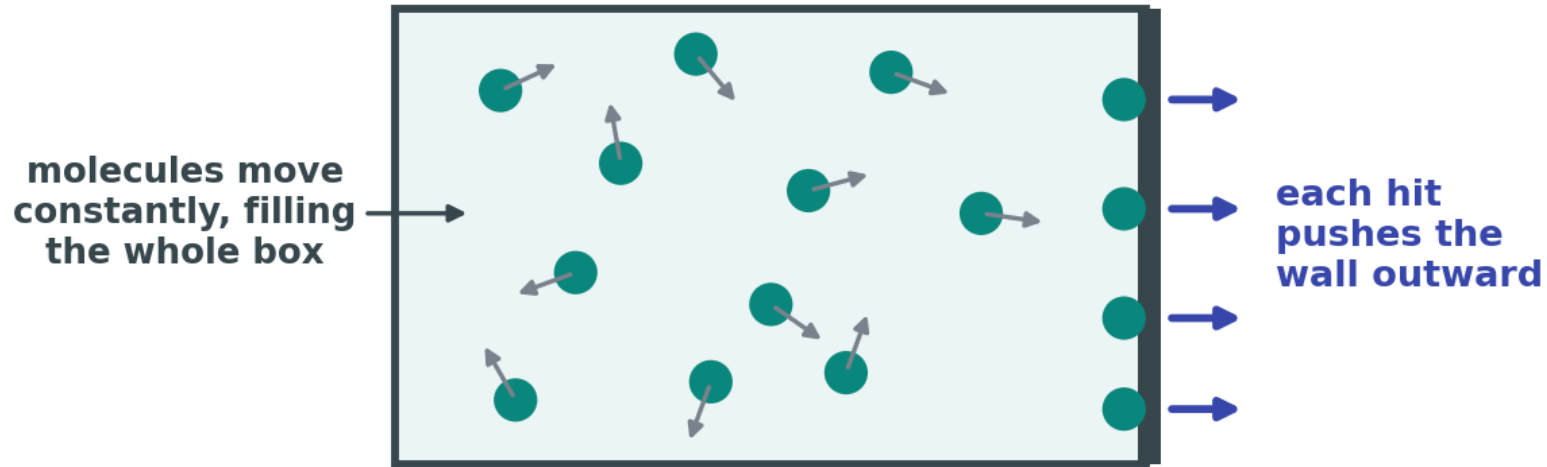
A bike tire pushes back



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

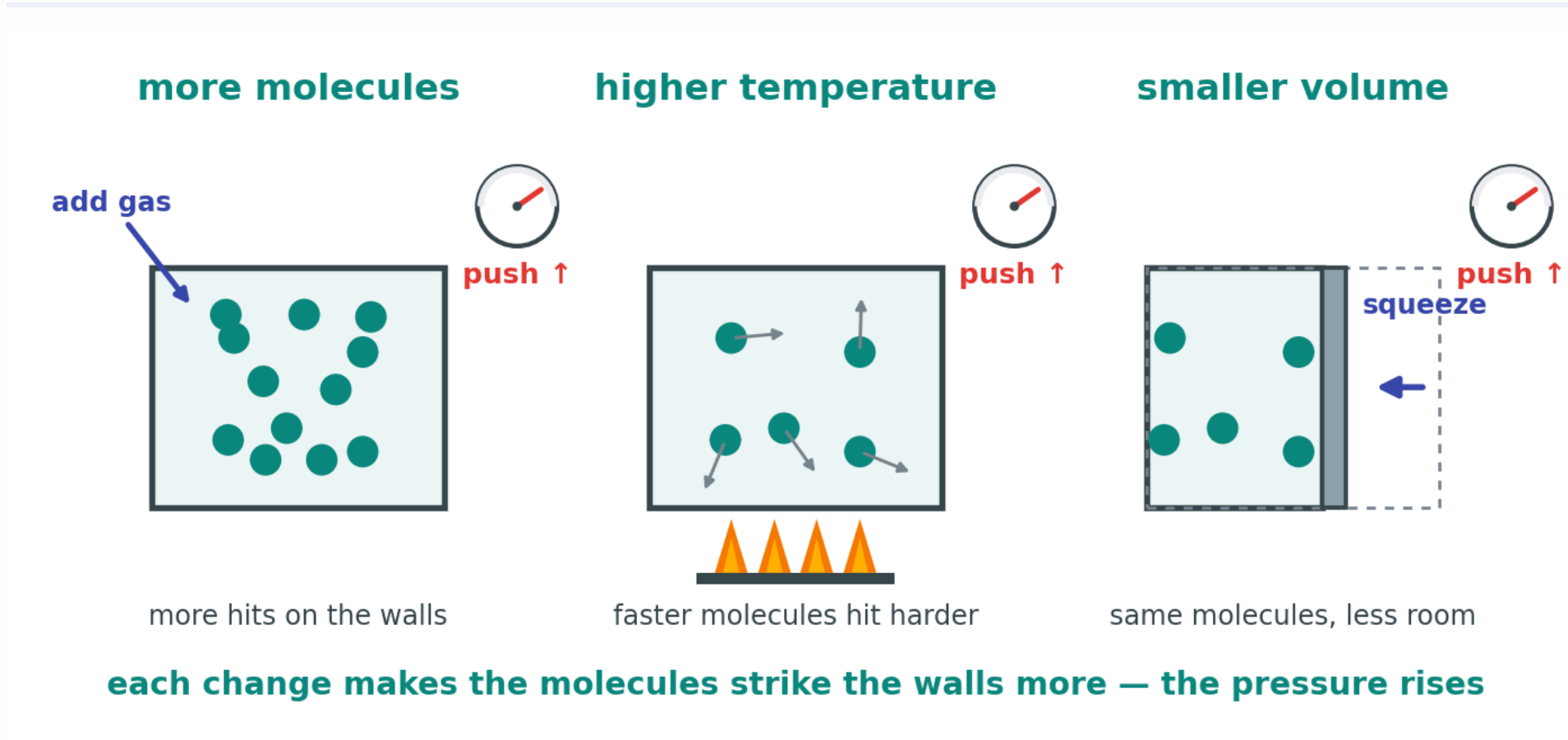
Air presses on every square inch

sea-level air: about 14.7 pounds on each square inch = 14.7 psi = 1 atm

psi = pounds per square inch: a force over an area, named in the unit itself

Sea-level air presses about 14.7 pounds onto every square inch of your skin, the weight of a bowling ball. The push comes from every side at once, so it goes unnoticed.

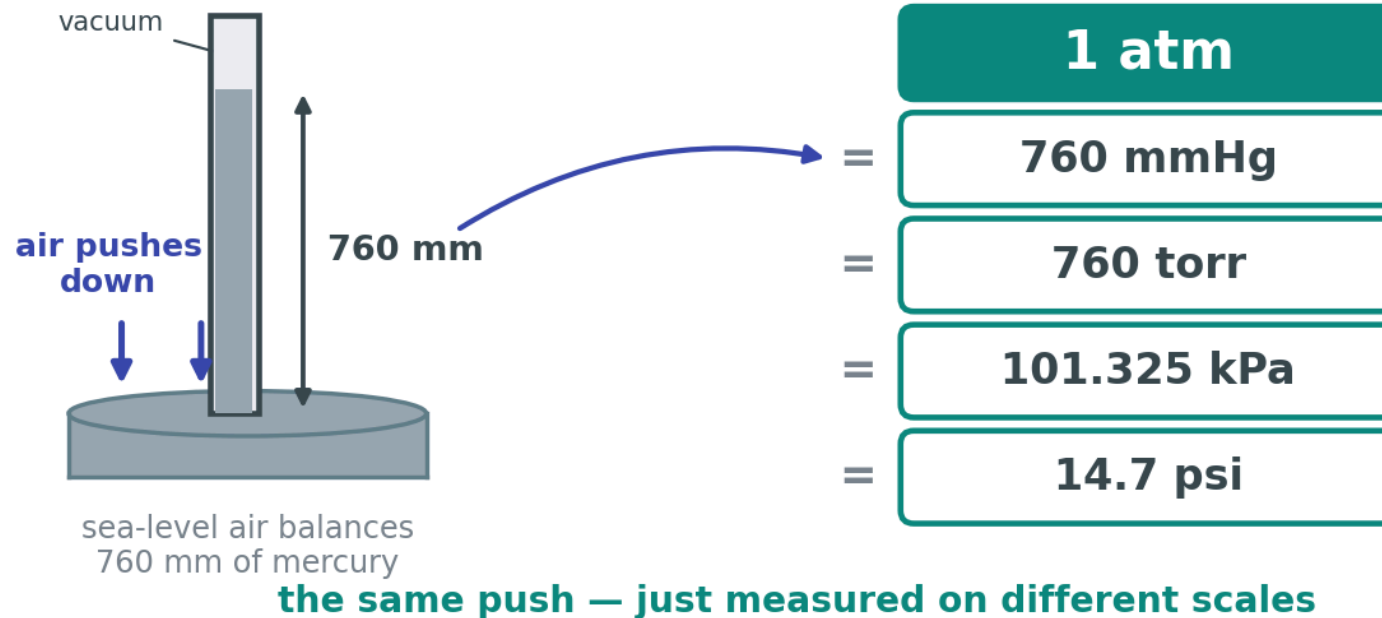
What raises the pressure



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 pascal and the kilopascal

$$1 \text{ atm} = 101,325 \text{ Pa} = 101.325 \text{ kPa}$$

1 kPa = 1000 Pa · a reading in Pa pairs with 101,325 · a reading in kPa pairs with 101.325

The SI unit of pressure is the pascal, Pa: one newton of force on each square meter. One pascal is a tiny push, so weather reports and tire gauges often use kilopascals.

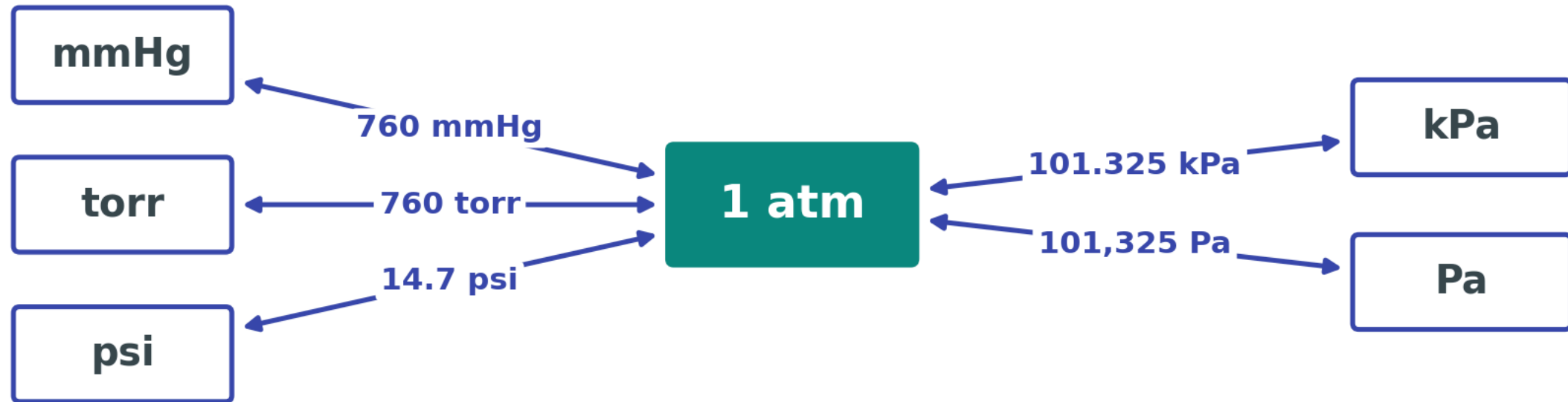


The method

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.

Every route runs through 1 atm

each spoke: the amount that equals 1 atm



Each unit has its own equality with 1 atm. Go from the given unit into atm, then out to the wanted unit. Each arrow on the map is one conversion factor.



Guided example: a mountain barometer

Step 1 • Write the given

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

given: 648 mmHg • wanted: atm

A barometer in a mountain town reads **648 mmHg**. Express the pressure in atmospheres.

Start from the equality that links the two units: 760 mmHg make 1 atm.

✓ Guided example: solution

1 atm = 760 mmHg
given: 648 mmHg · wanted: atm

Step 2 · Pick the factor that cancels its unit

Step 3 · Multiply; repeat until the wanted unit survives

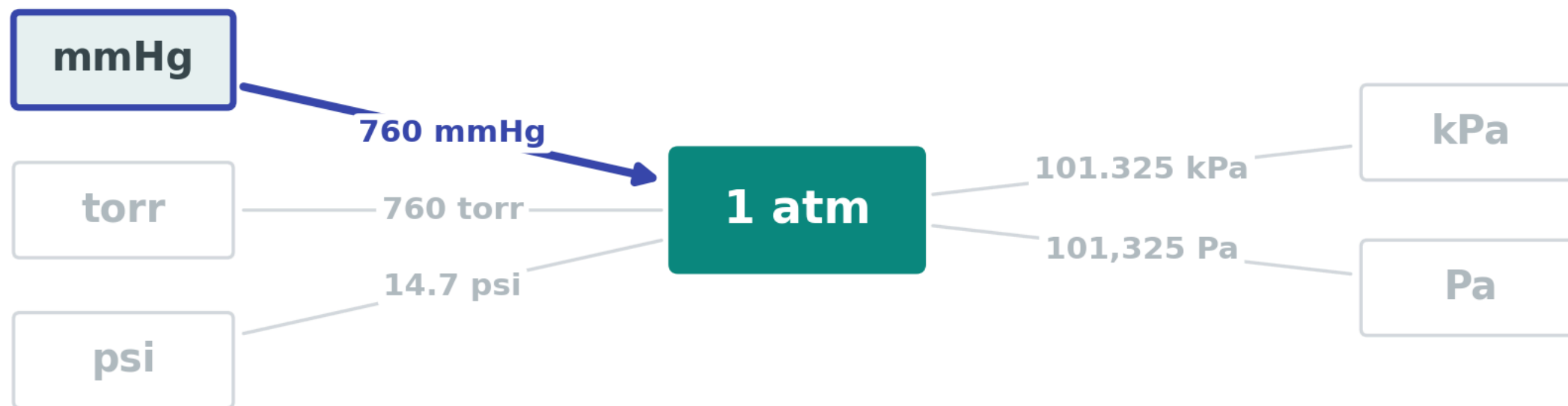
$$648 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.853 \text{ atm}$$

Step 4 · Sense-check

 **Sense check:** An atmosphere is a far larger unit than a millimeter of mercury, so the count shrinks: 648 → 0.853. Mountain air sits below 1 atm. ✓

✓ Guided example: the route on the map

$1 \text{ atm} = 760 \text{ mmHg}$
given: 648 mmHg · found: 0.853 atm



🧠 **Sense check:** mmHg to atm is one arrow on the map: one conversion factor. ✓



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

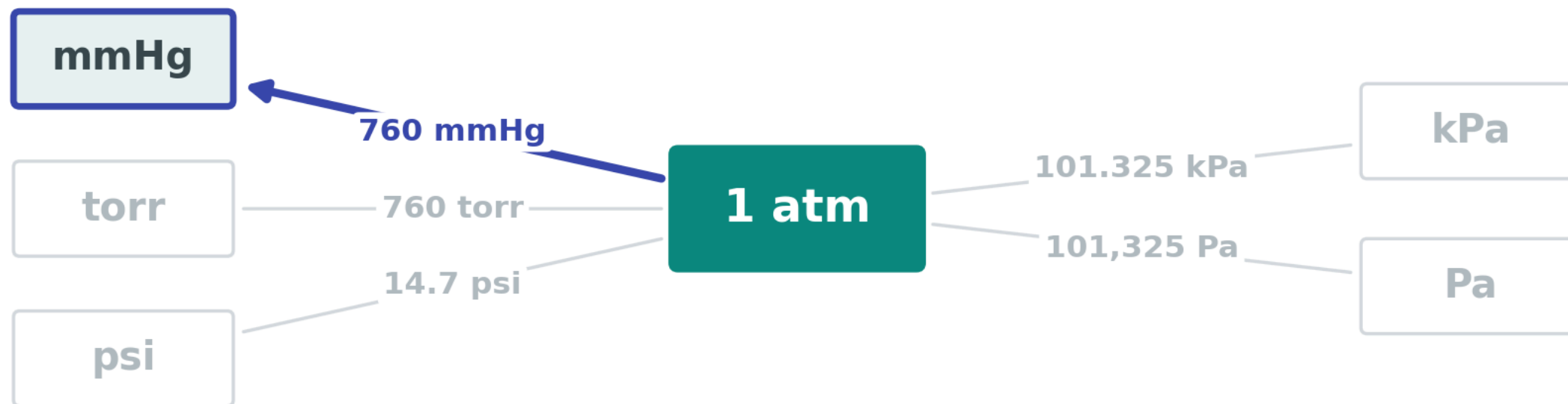
Step 4 · Sense-check

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

 **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

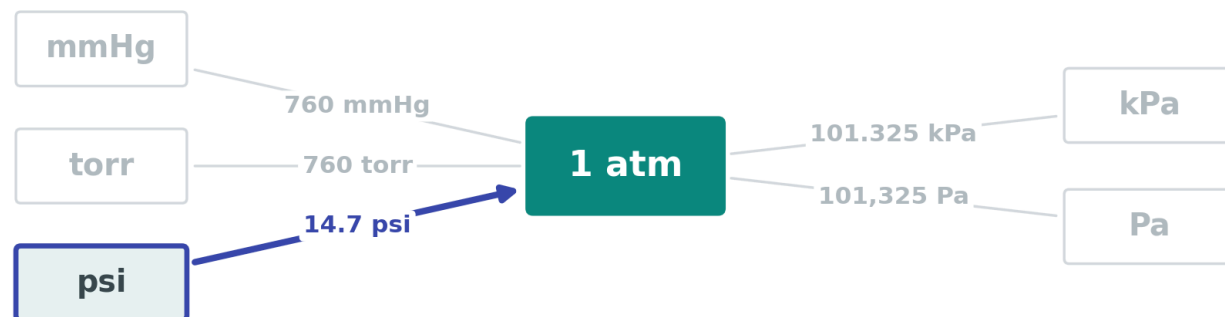
$$1 \text{ atm} = 14.7 \text{ psi}$$

given: 88.2 psi · wanted: atm

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

 **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. ✓

The route on the map



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

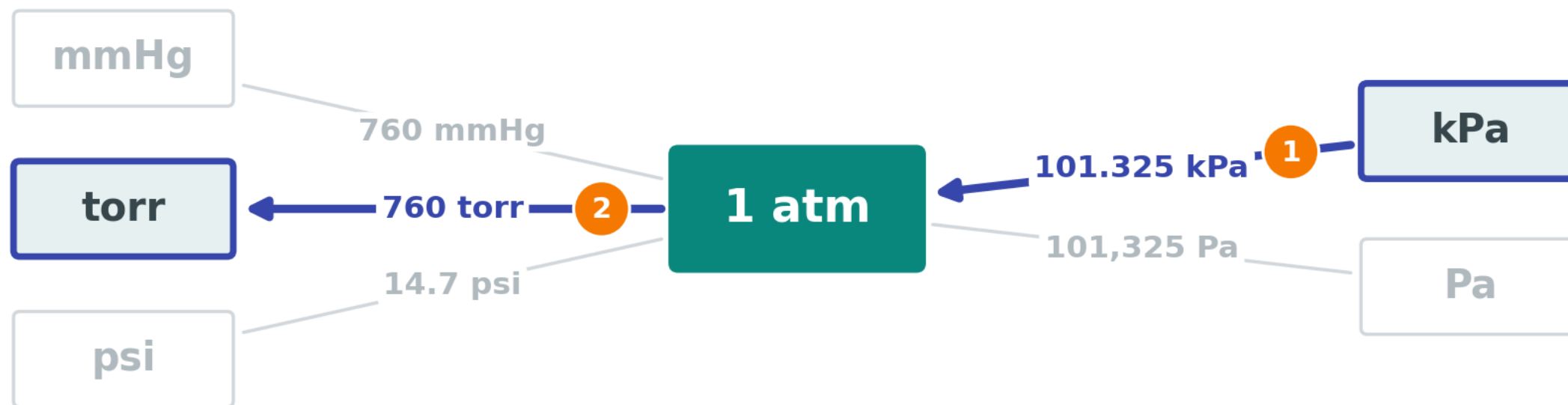
$$202.6 \text{ kPa} \times \frac{1 \text{ atm}}{101.325 \text{ kPa}} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 1520 \text{ torr}$$

🧠 **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: an air compressor

$$1 \text{ atm} = 14.7 \text{ psi} \cdot 1 \text{ atm} = 101.325 \text{ kPa}$$

given: 125 psi · wanted: kPa

A shop air compressor's tank gauge reads **125 psi**. What is this pressure in kPa?

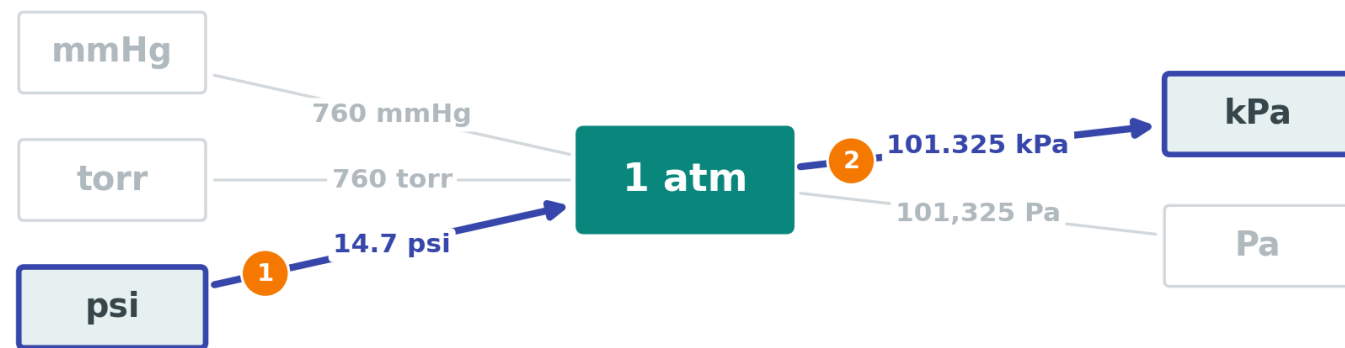
- A. 862
- B. 0.0839
- C. 6.46×10^3
- D. 8.50

✓ Practice 2 answer: A

1 atm = 14.7 psi · 1 atm = 101.325 kPa
 given: 125 psi · wanted: kPa · plan: psi → atm → kPa

$$125 \cancel{\text{psi}} \times \frac{1 \cancel{\text{atm}}}{14.7 \cancel{\text{psi}}} \times \frac{101.325 \text{ kPa}}{1 \cancel{\text{atm}}} = 862 \text{ kPa} \rightarrow \text{answer A}$$

🧠 **Sense check:** A kilopascal is a smaller unit than a psi, so the count grows: 125 → 862. ✓



Practice 3: a weather balloon

$$1 \text{ atm} = 101,325 \text{ Pa} \cdot 1 \text{ atm} = 760 \text{ mmHg}$$

given: $5.62 \times 10^4 \text{ Pa}$ · wanted: mmHg

High above the ground, a weather balloon's sensor reads $5.62 \times 10^4 \text{ Pa}$. What is this pressure in mmHg?

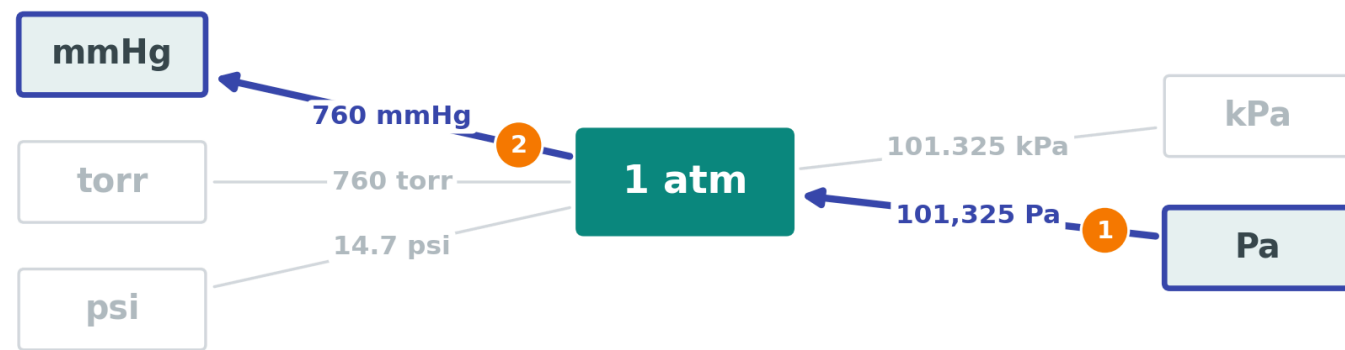
- A. 0.555
- B. 4.22×10^5
- C. 8.15
- D. 422
- E. 7.30×10^{-4}

✓ Practice 3 answer: D

1 atm = 101,325 Pa · 1 atm = 760 mmHg
 given: 5.62×10^4 Pa · wanted: mmHg · plan: Pa → atm → mmHg

$$5.62 \times 10^4 \text{ Pa} \times \frac{1 \text{ atm}}{101,325 \text{ Pa}} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} = 422 \text{ mmHg} \rightarrow \text{answer D}$$

🧠 **Sense check:** 5.62×10^4 Pa is a little over half of 101,325 Pa, so a little over half of 760 mmHg: 422. ✓



Practice 4: 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 4 answer: B

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

given: 758 torr $\rightarrow$ 162.0 kPa $\cdot$ two readings, two units: both go to atm before they subtract

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

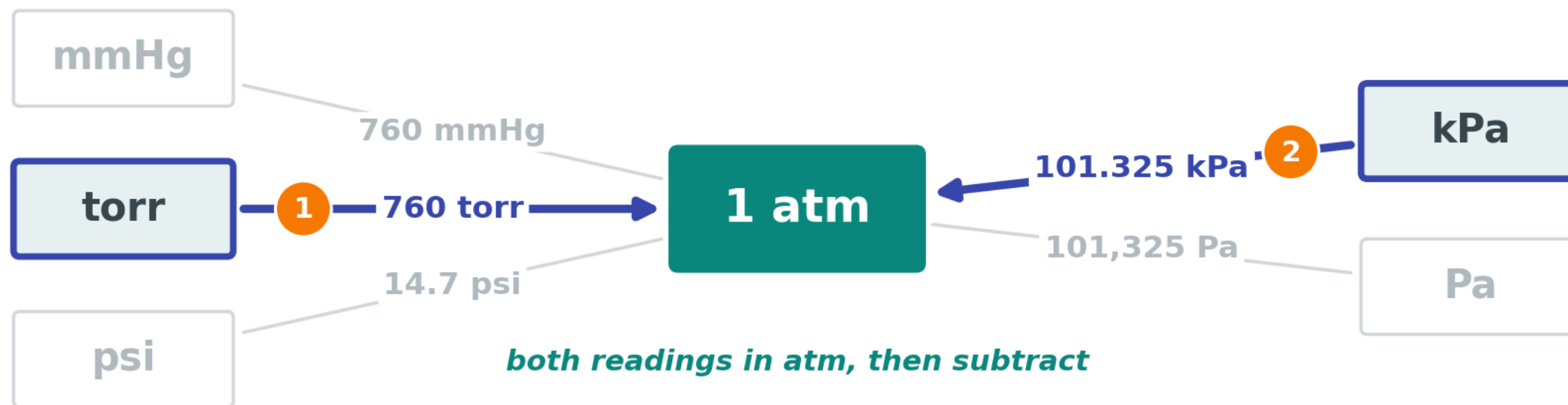
$$\text{rise} = 1.5988 \text{ atm} - 0.9974 \text{ atm} = \mathbf{0.601 \text{ atm} \rightarrow \text{answer B}}$$

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. ✓

✓ Practice 4: the route on the map

$1 \text{ atm} = 760 \text{ torr} \cdot 1 \text{ atm} = 101.325 \text{ kPa}$
 given: $758 \text{ torr} \rightarrow 162.0 \text{ kPa} \cdot$ found: a rise of 0.601 atm



🧠 **Sense check:** Two readings, two arrows, both into atm. Only then do the two numbers share a unit and subtract. ✓



Extra practice 1: an airliner cabin

$$1 \text{ atm} = 14.7 \text{ psi} \cdot 1 \text{ atm} = 760 \text{ mmHg}$$

given: 11.8 psi · wanted: mmHg

In flight, an airliner's cabin is held at **11.8 psi**. What is this pressure in mmHg?

- A. 0.803
- B. 81.3
- C. 1.06×10^{-3}
- D. 610.

✓ Extra practice 1 answer: D

$$1 \text{ atm} = 14.7 \text{ psi} \cdot 1 \text{ atm} = 760 \text{ mmHg}$$

given: 11.8 psi · wanted: mmHg · plan: psi → atm → mmHg

$$11.8 \text{ psi} \times \frac{1 \text{ atm}}{14.7 \text{ psi}} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} = 610. \text{ mmHg} \rightarrow \text{answer D}$$

 **Sense check:** Cabin air sits below 1 atm, so below 760 mmHg. A millimeter of mercury is a far smaller unit than a psi, so the count grows: 11.8 → 610. ✓



Extra practice 2: a fire extinguisher

$$1 \text{ atm} = 14.7 \text{ psi} \cdot 1 \text{ atm} = 101,325 \text{ Pa}$$

given: 195 psi · wanted: Pa

A fire extinguisher's gauge reads **195 psi**. What is this pressure in Pa?

- A. 13.3
- B. 1.34×10^3
- C. 1.34×10^6
- D. 1.31×10^{-4}
- E. 1.01×10^4

✓ Extra practice 2 answer: C

1 atm = 14.7 psi · 1 atm = 101,325 Pa
 given: 195 psi · wanted: Pa · plan: psi → atm → Pa

$$195 \cancel{\text{psi}} \times \frac{1 \cancel{\text{atm}}}{14.7 \cancel{\text{psi}}} \times \frac{101,325 \text{ Pa}}{1 \cancel{\text{atm}}} = 1.34 \times 10^6 \text{ Pa} \rightarrow \text{answer C}$$

 **Sense check:** A pascal is a tiny unit. Each atmosphere is 101,325 Pa, so 13.3 atm runs past a million pascals: 1.34×10^6 . ✓

Extra practice 3: a vacuum bell jar

$$1 \text{ atm} = 14.7 \text{ psi} \cdot 1 \text{ atm} = 760 \text{ mmHg}$$

given: 14.5 psi at the start · 85.0 mmHg after pumping · wanted: the fall in atm

A vacuum pump draws the air in a bell jar down from **14.5 psi** to **85.0 mmHg**. By how many atmospheres did the pressure fall?

- A. 0.875
- B. 0.112
- C. -0.875
- D. 0.148

✓ Extra practice 3 answer: A

$$1 \text{ atm} = 14.7 \text{ psi} \cdot 1 \text{ atm} = 760 \text{ mmHg}$$

given: 14.5 psi → 85.0 mmHg · two readings, two units: both go to atm before they subtract

$$14.5 \text{ psi} \times \frac{1 \text{ atm}}{14.7 \text{ psi}} = 0.9864 \text{ atm} \cdot 85.0 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.1118 \text{ atm}$$

$$\text{fall} = 0.9864 \text{ atm} - 0.1118 \text{ atm} = \mathbf{0.875 \text{ atm} \rightarrow \text{answer A}}$$

🧠 **Sense check:** From about 1 atm down to about a tenth of an atmosphere: a fall of a little under 0.9 atm. ✓



Check yourself

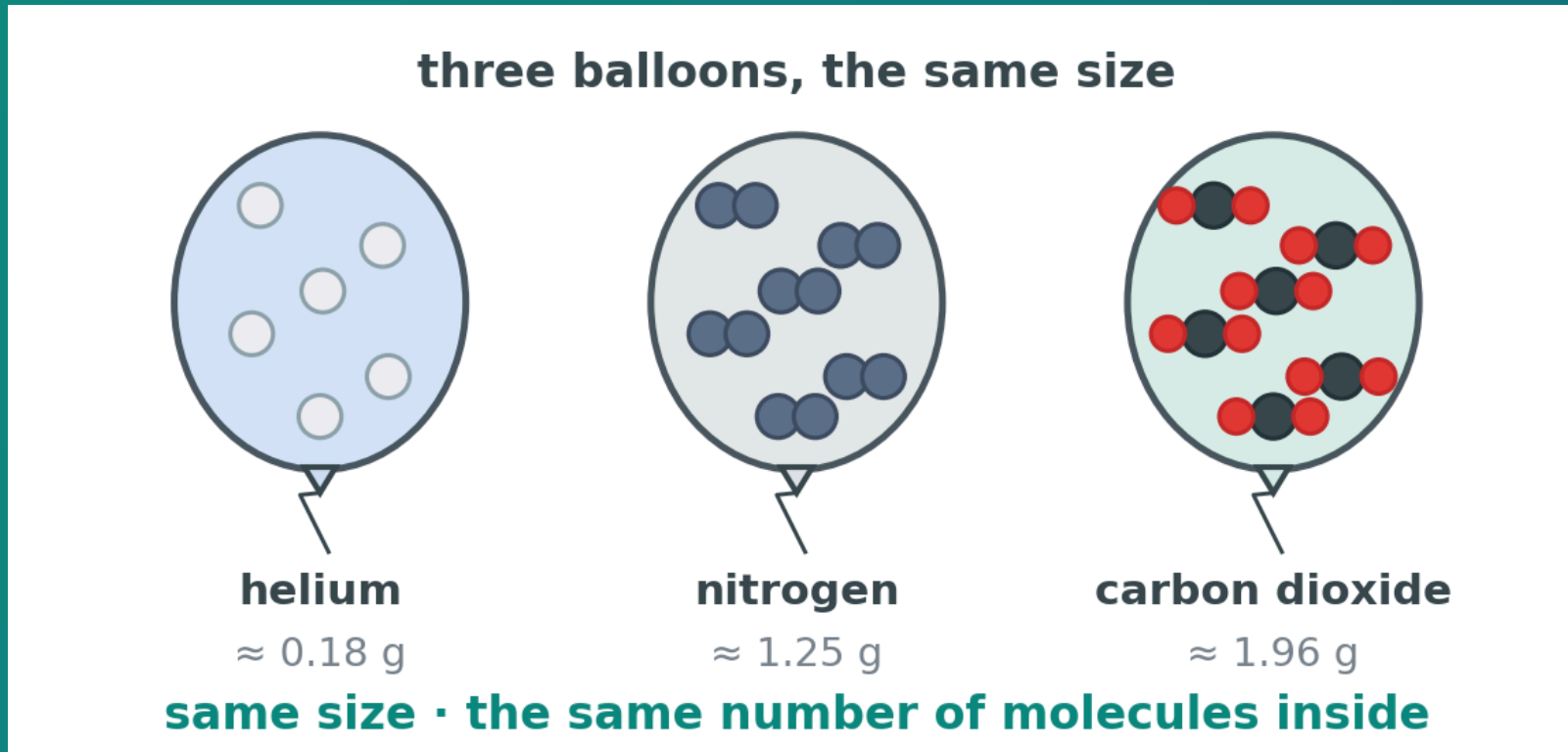
1. From **1 atm = 760 mmHg**, write both conversion factors. Which one converts a reading of 950 mmHg to atm?
2. A sealed rigid can is thrown on a fire. Do its molecules strike the walls harder or softer, and does the pressure rise or fall?

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

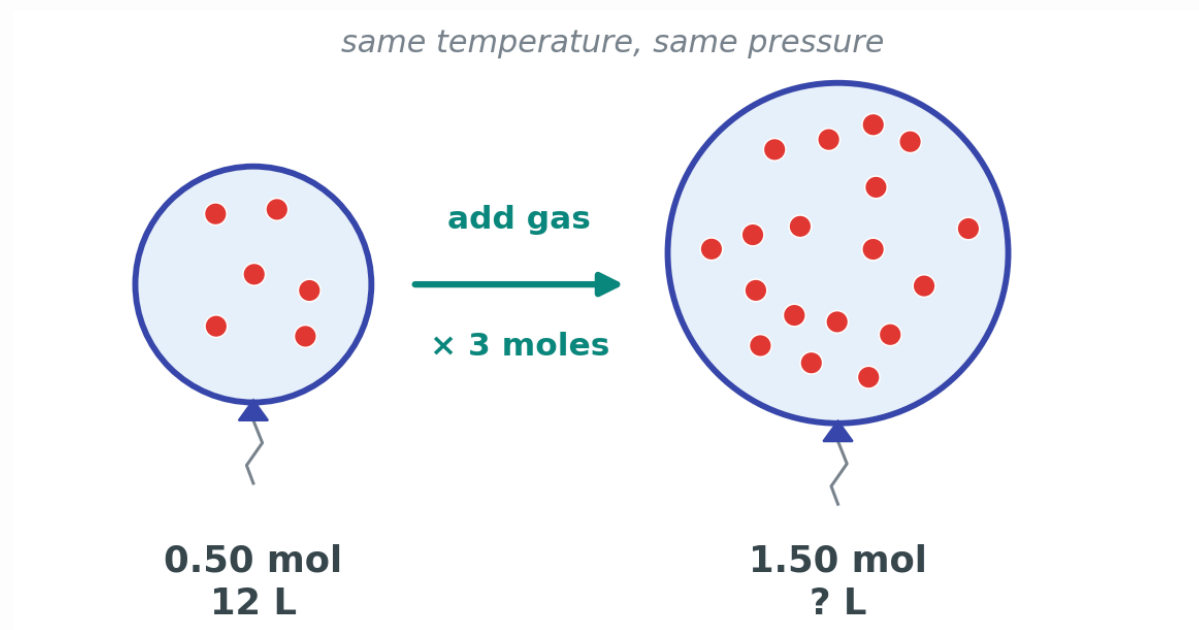
Use the molar volume of a gas at STP, 22.4 L/mol, to convert between moles and liters, and chain it with molar mass to go from grams of a gas to its volume, or to its density.

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.

More gas, a bigger balloon



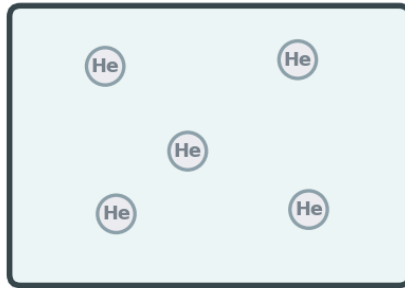
Blow more gas into a balloon and it grows. At the same temperature and pressure, the volume follows the moles. Predict the volume of the larger balloon.

$$1.50 \text{ mol} \div 0.50 \text{ mol} = 3 \cdot 3 \times 12 \text{ L} = 36 \text{ L}$$

three times the moles, three times the volume · which gas it is never entered

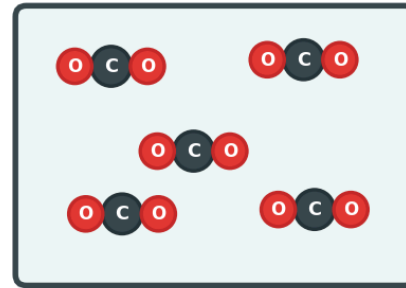
Equal volumes, equal counts

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.

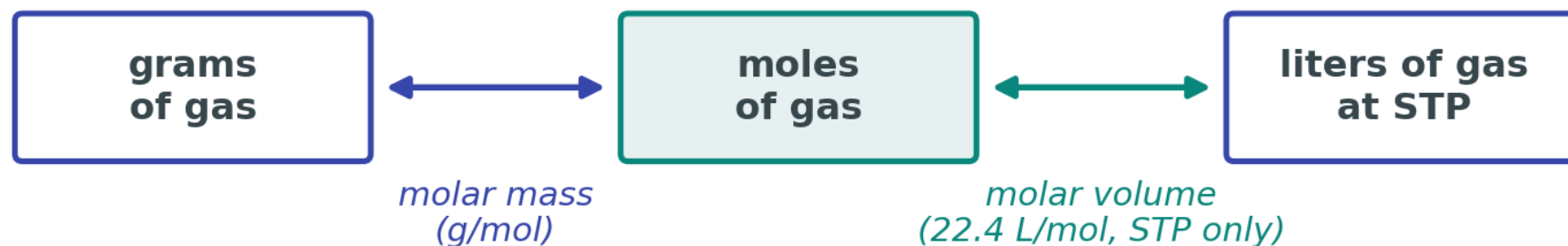
The molar volume is an equality

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



CO₂: 44.01 g = 1 mol = 22.4 L at STP

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.



Guided example: neon at 0 °C and 760 mmHg

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

given: 1.25 mol Ne · 0 °C · 760 mmHg · wanted: volume in L

A cylinder of neon for sign tubing holds **1.25 mol of Ne** at **0 °C and 760 mmHg**. What volume does the gas occupy?

Work the four method steps in order, starting with the conditions.

✓ Guided example: solution

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

given: 1.25 mol Ne · 0 °C · 760 mmHg · wanted: volume in L

Step 1 · Confirm STP

Step 2 · Map the route

Step 3 · Orient each factor

Step 4 · Multiply and check

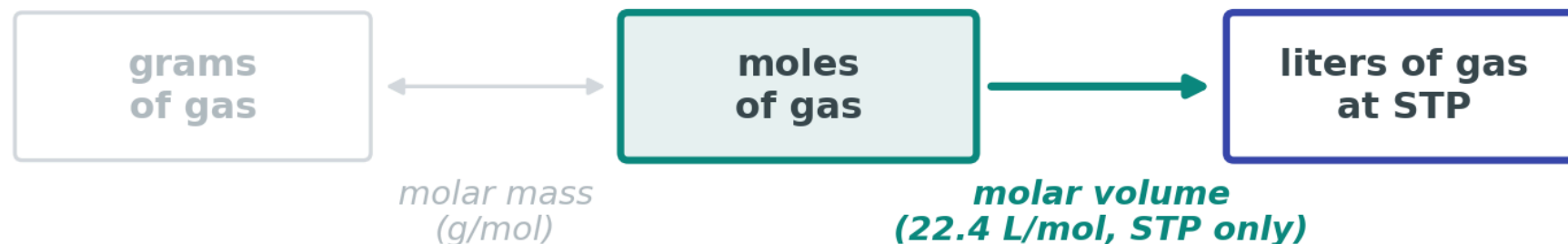
$$1.25 \text{ mol Ne} \times \frac{22.4 \text{ L}}{1 \text{ mol Ne}} = 28.0 \text{ L}$$

🧠 **Sense check:** One mole fills 22.4 L. A quarter mole more adds $22.4 \div 4 = 5.6 \text{ L}$: $22.4 + 5.6 = 28.0 \text{ L}$. ✓

✓ Guided example: the route on the map

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

given: 1.25 mol Ne at 0 °C and 760 mmHg · found: 28.0 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 1: liters to moles

Step 1 • Confirm STP

22.4 L = 1 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 1: 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

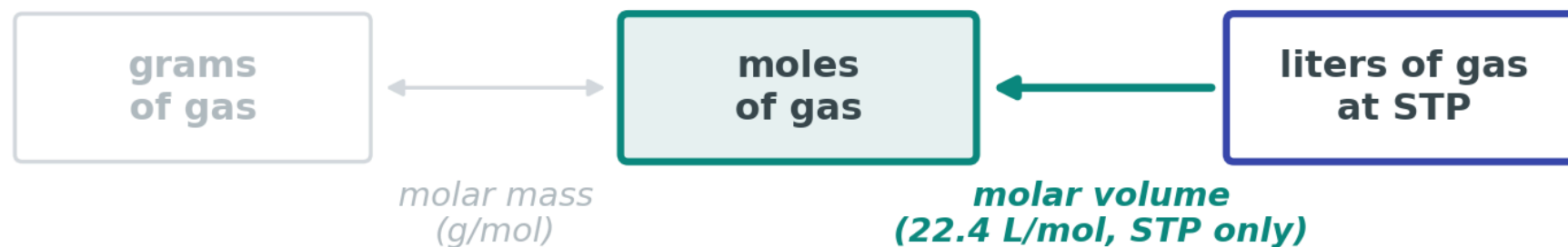
$$5.60 \text{ L} \times \frac{1 \text{ mol N}_2}{22.4 \text{ L}} = 0.250 \text{ mol N}_2$$

 **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 1: 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: 0.715 mol H₂ · wanted: L

A fuel-cell test stand draws **0.715 mol of hydrogen gas**, H₂, measured at STP. What volume, in liters, is that?

- A. 7.94
- B. 16.0
- C. 0.0319
- D. 1.44

✓ Practice 1 answer: B

22.4 L = 1 mol (at STP)

given: 0.715 mol H₂ · wanted: L · route: mol → L

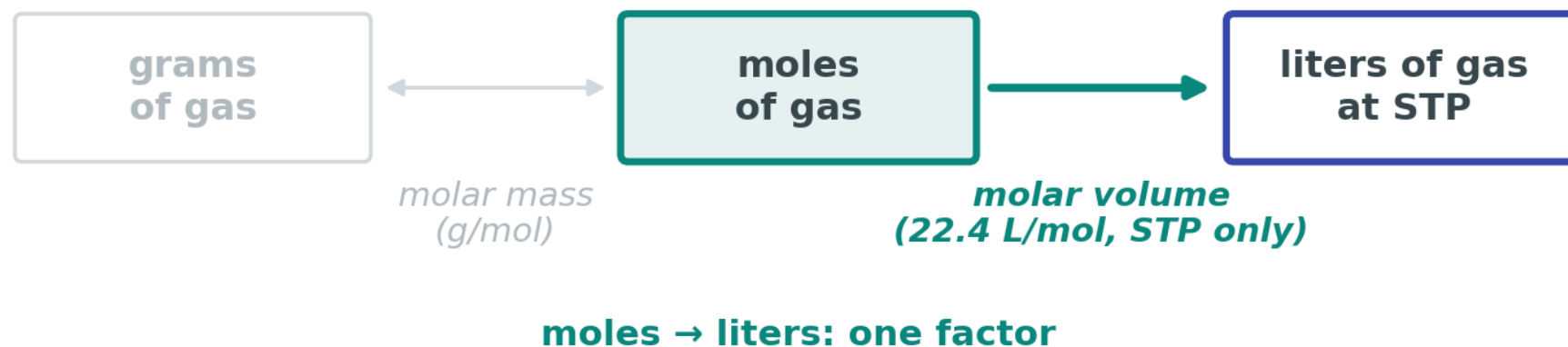
$$0.715 \cancel{\text{ mol H}_2} \times \frac{22.4 \text{ L}}{1 \cancel{\text{ mol H}_2}} = 16.0 \text{ L H}_2 \rightarrow \text{answer B}$$

A ran the grams chain on a mole amount: $0.715 / 2.016 \times 22.4 = 7.94$, as if 0.715 were grams. C flipped the molar volume: $0.715 / 22.4 = 0.0319$, in mol²/L. D reached for the molar mass: $0.715 \times 2.016 = 1.44$, a mass in grams.

🧠 **Sense check:** 0.715 mol is about 0.7 of a mole, so it fills about $0.7 \times 22.4 \approx 16 \text{ L}$. ✓

✓ Practice 1: the route on the map

22.4 L = 1 mol (at STP)
given: 0.715 mol H₂ · found: 16.0 L



🧠 **Sense check:** One arrow, moles to liters. A reached for the grams box, which this problem never visits. ✓



Practice 2

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 2 answer: D

22.4 L = 1 mol (at STP)

given: 3.36 L O₂ · wanted: mol O₂

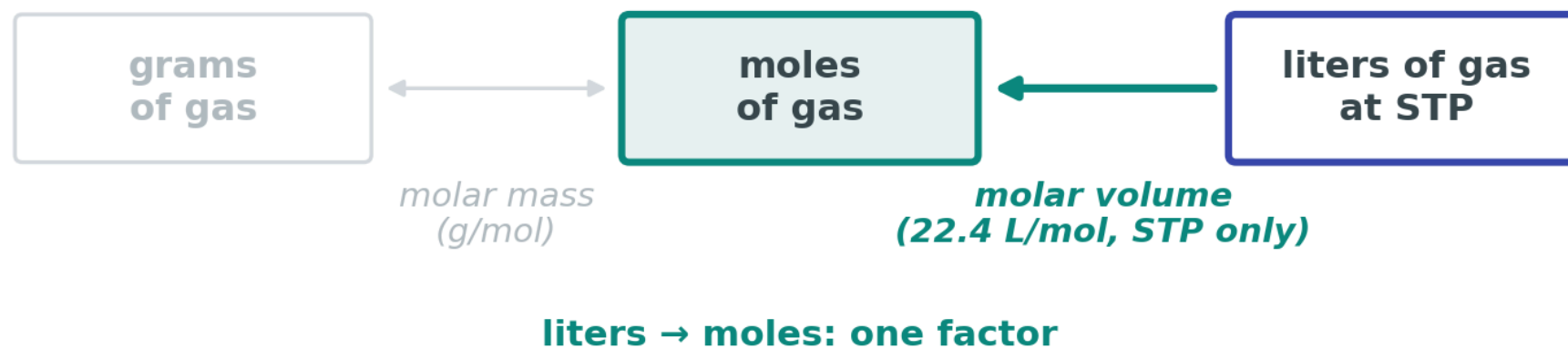
$$3.36 \cancel{\text{L O}_2} \times \frac{1 \text{ mol O}_2}{22.4 \cancel{\text{L O}_2}} = 0.150 \text{ mol O}_2 \rightarrow \text{answer D}$$

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. ✓

✓ Practice 2: the route on the map

22.4 L = 1 mol (at STP)
given: 3.36 L O₂ · found: 0.150 mol O₂



🧠 **Sense check:** Liters to moles is the molar-volume bridge alone. A and B crossed the molar-mass bridge, which leads to grams. ✓



Worked example 2: 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 2: 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

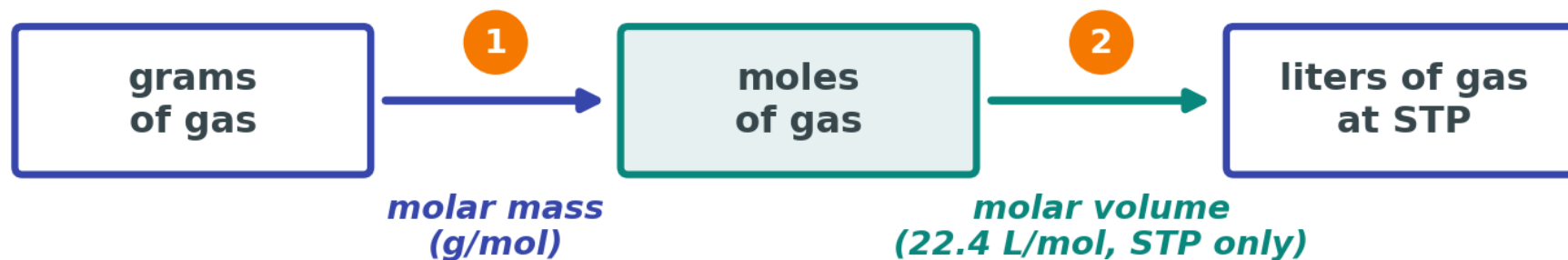
$$22.0 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{22.4 \text{ L}}{1 \text{ mol CO}_2} = 11.2 \text{ L}$$

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

✓ Worked example 2: 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. ✓



Worked example 3: liters to grams

Step 1 • Confirm STP

$$\text{CH}_4: 12.01 + 4(1.008) = 16.04 \text{ g/mol}$$

given: $1.00 \times 10^5 \text{ L CH}_4$ at STP • wanted: mol CH_4 , then g CH_4 • molar volume 22.4 L/mol

A natural-gas storage tank holds **$1.00 \times 10^5 \text{ L of methane}$** , CH_4 , at STP. How many moles of methane are in the tank? How many grams is that?

Map the route before writing any factor.

✓ Worked example 3: solution

CH₄: 16.04 g/mol · 22.4 L = 1 mol at STP
given: 1.00 × 10⁵ L CH₄ · wanted: mol, then g

Step 2 · Map the route

Step 3 · Orient each factor

Step 4 · Multiply and check

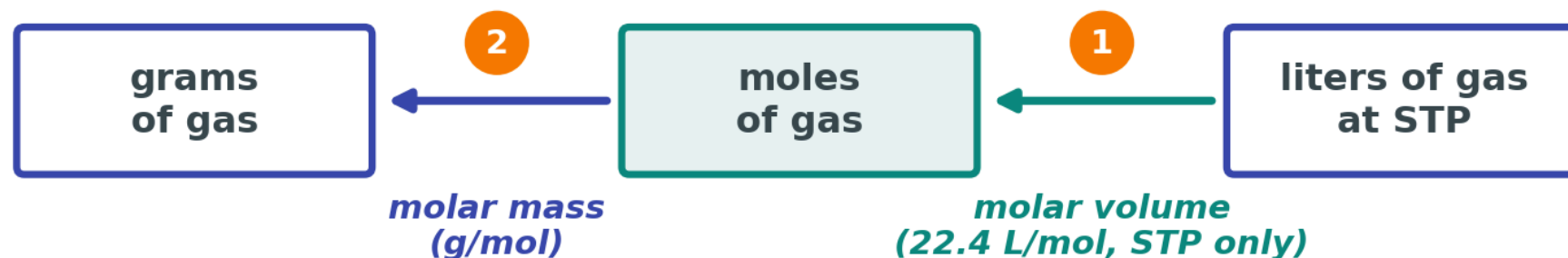
$$1.00 \times 10^5 \cancel{\text{L CH}_4} \times \frac{1 \text{ mol CH}_4}{22.4 \cancel{\text{L CH}_4}} = 4.46 \times 10^3 \text{ mol CH}_4$$

$$1.00 \times 10^5 \cancel{\text{L CH}_4} \times \frac{1 \cancel{\text{mol CH}_4}}{22.4 \cancel{\text{L CH}_4}} \times \frac{16.04 \text{ g CH}_4}{1 \cancel{\text{mol CH}_4}} = 7.16 \times 10^4 \text{ g CH}_4$$

🧠 **Sense check:** One mole is 16.04 g in 22.4 L, about 0.716 g per liter. 1.00 × 10⁵ L × 0.7 g/L ≈ 7 × 10⁴ g. ✓

✓ Worked example 3: the route on the map

CH₄: 16.04 g/mol · 22.4 L = 1 mol at STP
 given: 1.00 × 10⁵ L CH₄ · found: 4.46 × 10³ mol, 7.16 × 10⁴ g



liters → moles → grams: two factors

Sense check: Liters in, grams out: the molar volume reaches moles, then the molar mass reaches grams. The moles are the stop between. ✓



Practice 3

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 3 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: $\text{L} \rightarrow \text{mol} \rightarrow \text{g}$

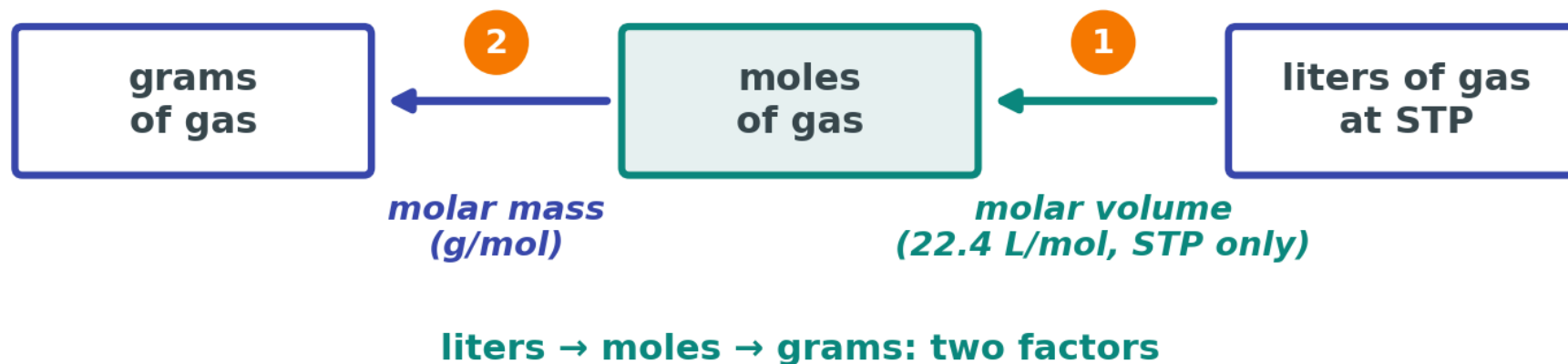
$$6.72 \text{ L SO}_2 \times \frac{1 \text{ mol SO}_2}{22.4 \text{ L SO}_2} \times \frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} = 19.2 \text{ g SO}_2 \rightarrow \text{answer A}$$

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 . ✓

✓ Practice 3: the route on the map

SO₂: 64.07 g/mol · 22.4 L = 1 mol at STP
 given: 6.72 L SO₂ · found: 0.300 mol, 19.2 g SO₂



🧠 **Sense check:** Both bridges, in order: liters to moles, then moles to grams. B stopped after the first. ✓



Worked example 4: density of a gas at STP

Step 1 • Confirm STP

O_2 : $2(16.00) = 32.00 \text{ g/mol} \cdot 22.4 \text{ L} = 1 \text{ mol at STP}$

density = mass $\div$ volume • wanted: density of O_2 at STP, in g/L

A hospital supply room stores cylinders of oxygen. What is the density of oxygen gas, O_2 , at STP?

Density is grams per liter. Set it up from one mole.

✓ Worked example 4: solution

O_2 : $32.00 \text{ g/mol} \cdot 22.4 \text{ L} = 1 \text{ mol}$ at STP

wanted: $\text{g/L} \cdot \text{density} = \text{mass} \div \text{volume}$

Step 2 · Map the route

Step 3 · Orient each factor

Step 4 · Multiply and check

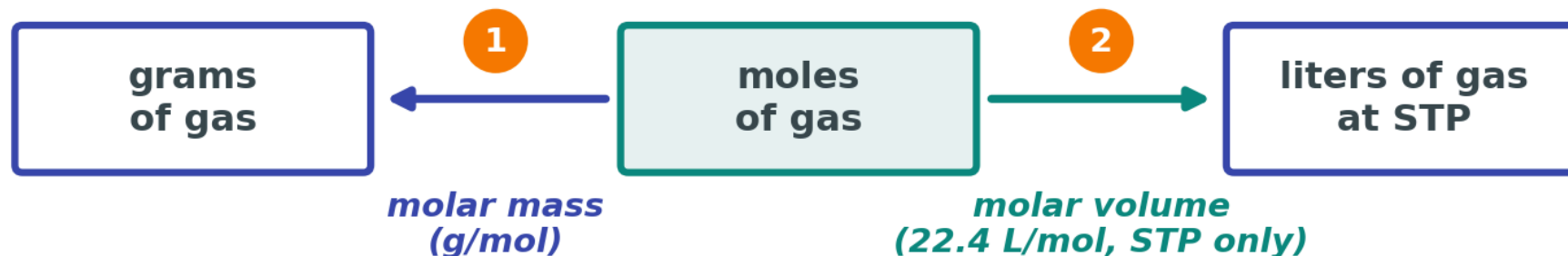
$$\frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \times \frac{1 \text{ mol O}_2}{22.4 \text{ L}} = 1.43 \text{ g/L}$$

 **Sense check:** Gas densities are small: about a gram per liter, not per milliliter. Every gas fills 22.4 L per mole, so the heavier the molecule, the denser the gas. ✓

✓ Worked example 4: the route on the map

$$\text{O}_2: 32.00 \text{ g/mol} \cdot 22.4 \text{ L} = 1 \text{ mol at STP}$$

found: 1.43 g/L



one mole on both bridges: density = grams ÷ liters

 **Sense check:** Both bridges leave from one mole: the molar mass gives its grams, the molar volume its liters. Density divides the two. ✓



Practice 4

22.4 L = 1 mol at STP

given: NO₂ gas at STP · wanted: density in g/L

Smog over a city takes its brown color from nitrogen dioxide, NO₂. What is the density of NO₂ gas at STP, in g/L?

- A. 1.34
- B. 0.487
- C. 2.05
- D. 46.0

✓ Practice 4 answer: C

NO_2 : $14.01 + 2(16.00) = 46.01 \text{ g/mol} \cdot 22.4 \text{ L} = 1 \text{ mol at STP}$

wanted: g/L · route: one mole on both bridges

$$\frac{46.01 \text{ g NO}_2}{1 \text{ mol NO}_2} \times \frac{1 \text{ mol NO}_2}{22.4 \text{ L}} = 2.05 \text{ g/L} \rightarrow \text{answer C}$$

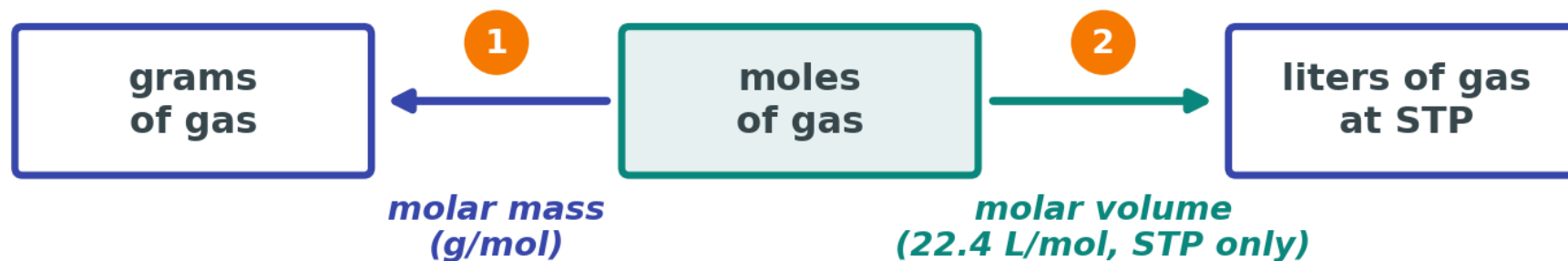
A counted one oxygen: NO is $14.01 + 16.00 = 30.01 \text{ g/mol}$, and $30.01 / 22.4 = 1.34$. B flipped the density: $22.4 / 46.01 = 0.487$, in L/g. D stopped at the molar mass: 46.0 g is the mass of a mole, not of a liter.

🧠 **Sense check:** NO_2 , at 46.01 g/mol, is heavier than O_2 , at 32.00 g/mol, so it is the denser gas: $2.05 \text{ g/L} > 1.43 \text{ g/L}$. ✓

✓ Practice 4: the route on the map

$$\text{NO}_2: 46.01 \text{ g/mol} \cdot 22.4 \text{ L} = 1 \text{ mol at STP}$$

found: 2.05 g/L



one mole on both bridges: density = grams ÷ liters

🧠 **Sense check:** One mole on both bridges: 46.01 g over 22.4 L. D stopped after the first bridge. ✓



Practice 5

22.4 L = 1 mol at STP

given: bulb volume · empty and filled masses · wanted: molar mass in g/mol

A student evacuates a **350. mL** glass bulb and weighs it: **35.462 g**. Filled with an unknown gas at STP, the bulb weighs **36.570 g**. What is the molar mass of the gas, in g/mol?

- A. 3.17
- B. 2.34×10^3
- C. 24.8
- D. 7.08
- E. 70.9

✓ Practice 5 answer: E

gas: $36.570 - 35.462 = 1.108 \text{ g} \cdot 350. \text{ mL} = 0.350 \text{ L at STP}$

route: $\text{L} \rightarrow \text{mol}$, then $\text{g} \div \text{mol} = \text{molar mass}$

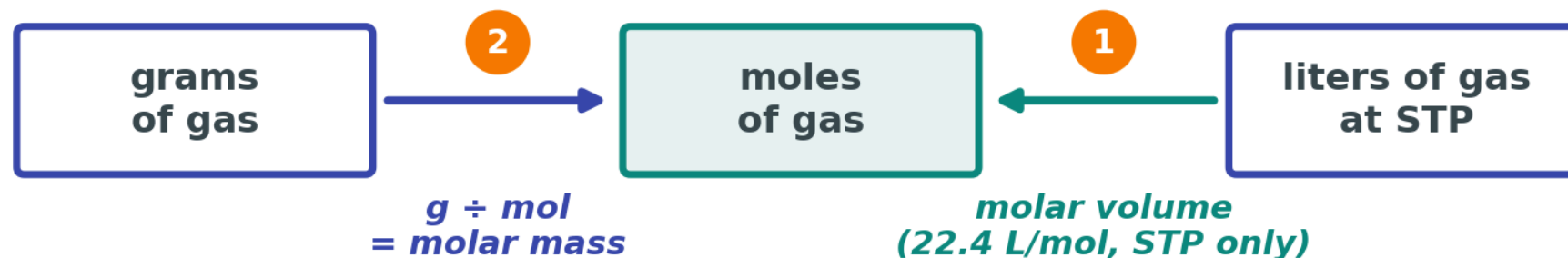
$$0.350 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.01563 \text{ mol, then } \frac{1.108 \text{ g}}{0.01563 \text{ mol}} = \mathbf{70.9 \text{ g/mol} \rightarrow \text{answer E}}$$

A stopped at the density: $1.108 / 0.350 = 3.17 \text{ g/L}$. B used the filled bulb, glass and all: $36.570 / 0.01563 = 2.34 \times 10^3$. C skipped the volume, as if the bulb held 1 L: $1.108 \times 22.4 = 24.8$. D flipped the density: $0.350 / 1.108 \times 22.4 = 7.08$.

🧠 **Sense check:** 70.9 g/mol matches Cl_2 , $2(35.45) = 70.90 \text{ g/mol}$. ✓

✓ Practice 5: the route on the map

gas: 1.108 g · 0.350 L at STP
found: 0.01563 mol · 70.9 g/mol



liters → moles, then grams ÷ moles = molar mass

🧠 **Sense check:** The molar volume turns the bulb's liters into moles. Grams over moles is the molar mass. ✓



Extra practice 1

22.4 L = 1 mol at STP

given: 25.5 g NH_3 at STP · wanted: L

A refrigeration line leaks **25.5 g of ammonia**, NH_3 . What volume, in liters, does that gas fill at STP?

- A. 571
- B. 33.5
- C. 0.0668
- D. 1.50

✓ Extra practice 1 answer: B

NH_3 : $14.01 + 3(1.008) = 17.03 \text{ g/mol} \cdot 22.4 \text{ L} = 1 \text{ mol at STP}$

given: 25.5 g NH_3 · wanted: L · route: $\text{g} \rightarrow \text{mol} \rightarrow \text{L}$

$$25.5 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} \times \frac{22.4 \text{ L}}{1 \text{ mol NH}_3} = 33.5 \text{ L NH}_3 \rightarrow \text{answer B}$$

A used the grams as moles: $25.5 \times 22.4 = 571$, with no molar mass. C flipped the molar volume: $1.50 / 22.4 = 0.0668$, in mol^2/L . D stopped at moles: $25.5 / 17.03 = 1.50 \text{ mol}$, an amount, not a volume.

🧠 **Sense check:** 25.5 g is just under 1.50 mol of NH_3 , so the gas fills just under $1.5 \times 22.4 = 33.6 \text{ L}$. ✓



Extra practice 2

22.4 L = 1 mol at STP

given: 1.78 g/L · 0 °C · 101.325 kPa · wanted: g/mol

An unknown gas has a density of **1.78 g/L** at **0 °C** and **101.325 kPa**. What is the molar mass of the gas, in g/mol?

- A. 12.6
- B. 0.0795
- C. 1.78
- D. 39.9

✓ Extra practice 2 answer: D

101.325 kPa = 1 atm, so 0 °C and 101.325 kPa is STP · 22.4 L = 1 mol

given: 1.78 g/L · wanted: g/mol · route: g per L → g per mol

$$\frac{1.78 \text{ g}}{1 \text{ L}} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 39.9 \text{ g/mol} \rightarrow \text{answer D}$$

A divided the molar volume by the density: $22.4 / 1.78 = 12.6$, units that never reduce to g/mol. B flipped the molar volume: $1.78 / 22.4 = 0.0795$, in g·mol/L². C took the density for the molar mass: 1.78 g is the mass of one liter, not of one mole.

 **Sense check:** The gas is denser than O₂, 1.43 g/L, so its molar mass sits above 32.00. 39.9 g/mol matches Ar, 39.95 g/mol. ✓



Extra practice 3

22.4 L = 1 mol at STP

given: rigid 7.84 L tank · 0.120 mol N₂ at 0 °C · wanted: mol N₂ to add

A rigid **7.84 L** tank at 0 °C holds **0.120 mol** of nitrogen, N₂. How many more moles of N₂ must be added at 0 °C to bring the pressure to 1 atm?

- A. 0.230
- B. 0.470
- C. 175
- D. 0.350

✓ Extra practice 3 answer: A

0 °C and 1 atm is STP · 22.4 L = 1 mol

given: 7.84 L tank · 0.120 mol N₂ present · route: L → mol at 1 atm, then subtract

At 0 °C and 1 atm the tank is at STP. Full, it holds one mole per 22.4 L.

$$7.84 \text{ L} \times \frac{1 \text{ mol N}_2}{22.4 \text{ L}} = 0.350 \text{ mol N}_2 \text{ at 1 atm} \rightarrow 0.350 - 0.120 = \mathbf{0.230 \text{ mol N}_2 \rightarrow \text{answer A}}$$

B added the gas already present: $0.350 + 0.120 = 0.470$. C flipped the molar volume: $7.84 \times 22.4 - 0.120 = 175$, a number that is not moles. D stopped at the full tank: 0.350 mol, but 0.120 mol is already inside.

 **Sense check:** The tank holds 0.120 of the 0.350 mol it needs, about a third. Roughly two-thirds is missing: 0.230 mol. ✓

Check yourself

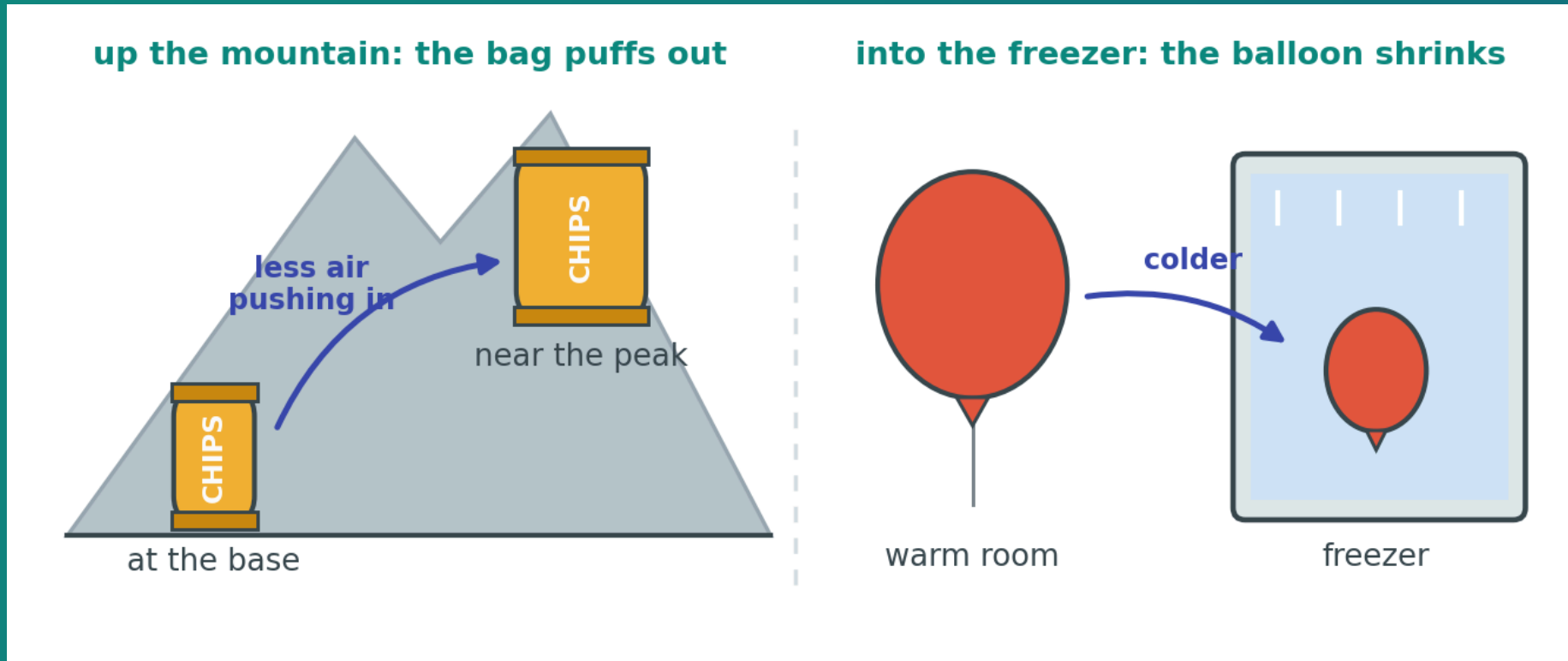
1. From **22.4 L = 1 mol at STP**, write both conversion factors. Which one converts 0.300 mol of a gas to liters, and which converts 44.8 L of a gas to moles?
2. One flask holds 1.00 mol of helium; another holds 1.00 mol of carbon dioxide. Both are at STP. Which flask has the greater volume? Which holds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

Pressure and volume trade off

$P \text{ (atm, mmHg, kPa)} \cdot V \text{ (L, mL)} \cdot T \text{ (K)} \cdot n \text{ (mol)}$
four quantities describe a gas · a sealed container fixes n

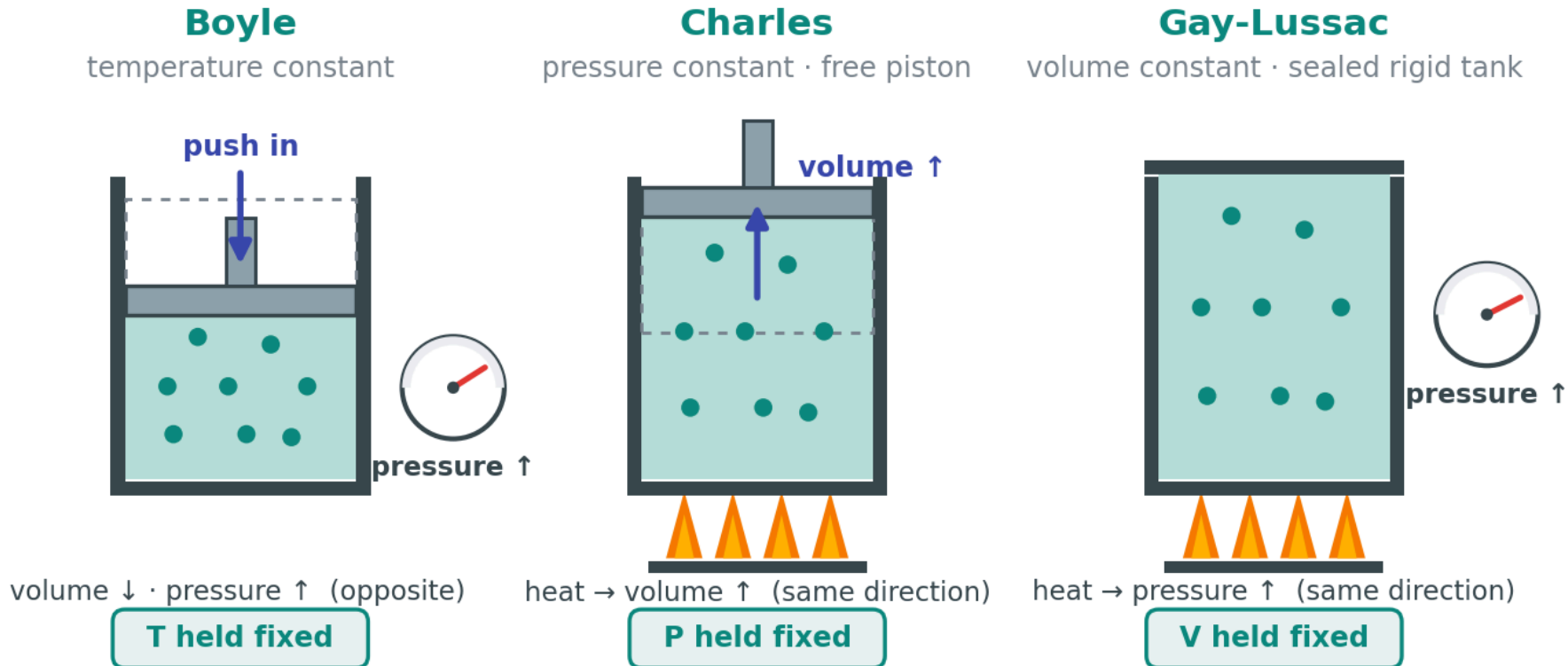
Squeeze a gas at constant temperature. Its particles hit the walls more often, so pressure rises. A 4.0 L balloon at 1.0 atm is taken underwater:

1.0 atm $\rightarrow$ 2.0 atm: pressure $\times 2 \rightarrow$ volume $\div 2 = \mathbf{2.0 \text{ L}}$

Deeper, the pressure reaches 4.0 atm. Find the volume.

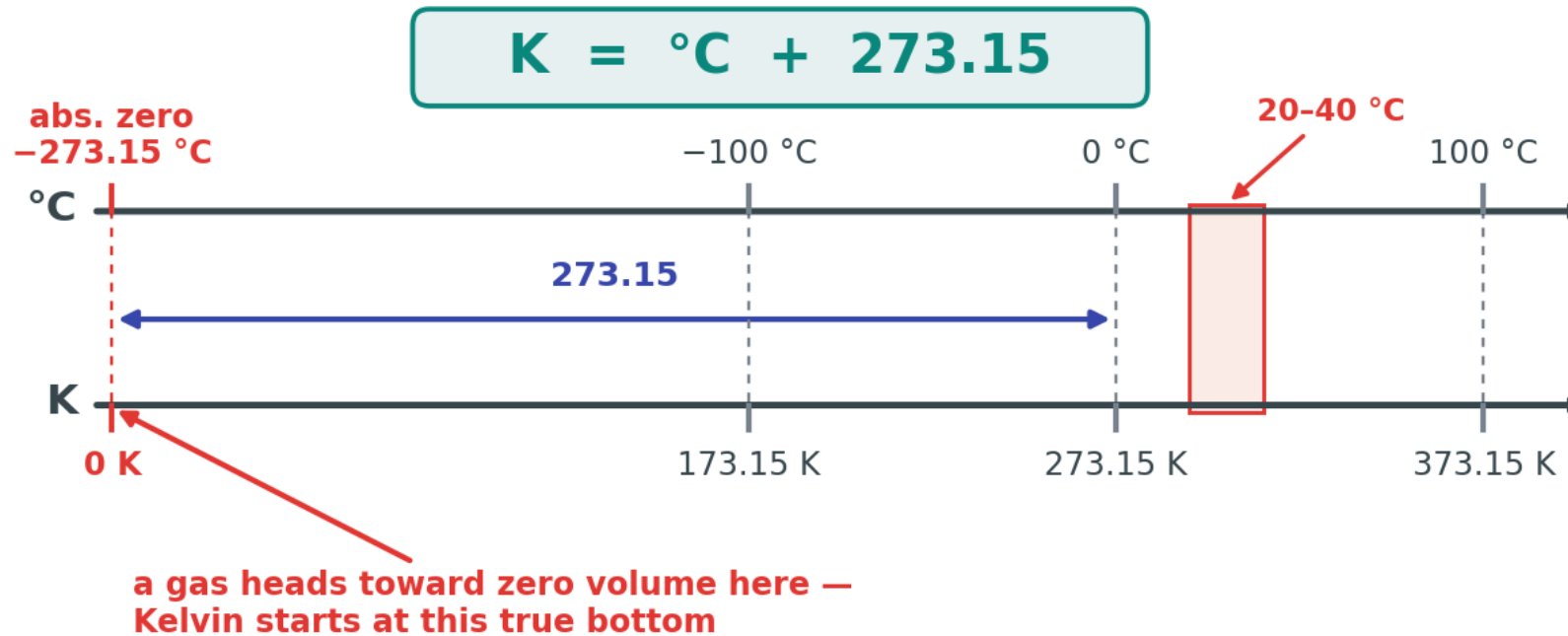
1.0 atm $\rightarrow$ 4.0 atm: pressure $\times 4 \rightarrow$ volume $\div 4 = \mathbf{1.0 \text{ L}}$

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

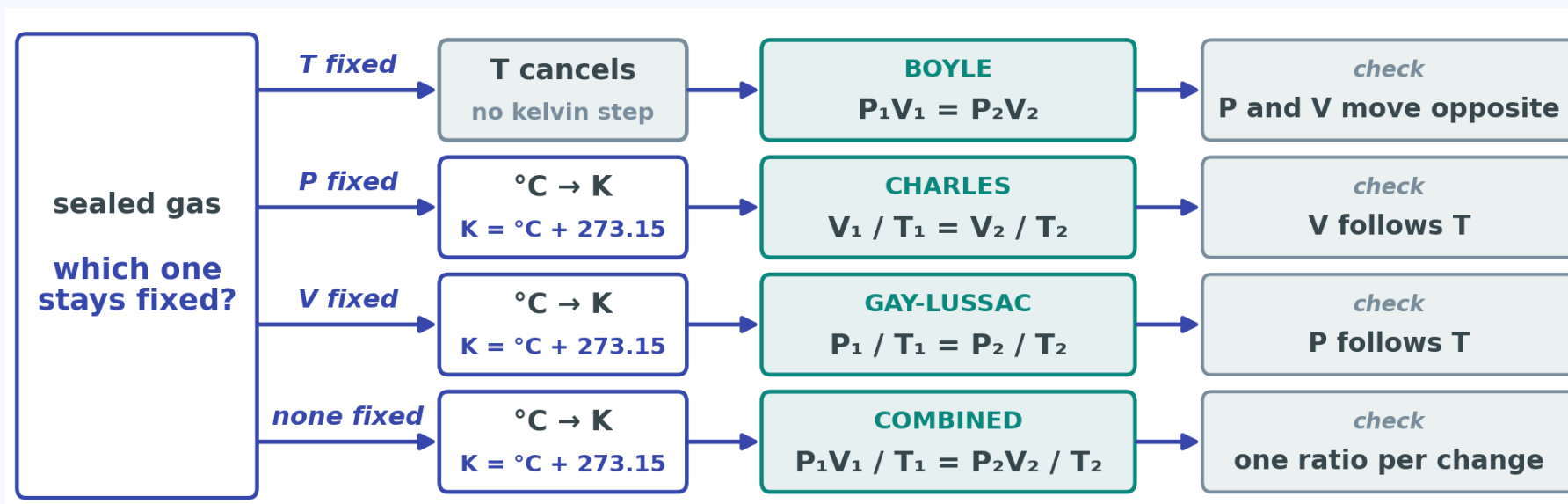


20 °C and 40 °C are not double: as 293.15 K and 313.15 K they sit about 7% apart. A gas responds to the Kelvin ratio.

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?





Guided example: a steel tank in the sun

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

given: $P_1 = 655 \text{ kPa}$ · $12.0 \text{ }^\circ\text{C} \rightarrow 42.0 \text{ }^\circ\text{C}$ · rigid, sealed tank · wanted: P_2

A rigid, sealed steel tank holds air at **655 kPa** on a cool morning, **12.0 °C**. By afternoon the sun has warmed it to **42.0 °C**. Find the new pressure.

Sort the quantities first: which ones change, and which one stays fixed?

✓ Guided example: pick the law

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

given: $P_1 = 655 \text{ kPa} \cdot 12.0 \text{ }^\circ\text{C} \rightarrow 42.0 \text{ }^\circ\text{C} \cdot \text{rigid, sealed tank} \cdot \text{wanted: } P_2$

Step 1 • Identify given and wanted

changes: P and $T \cdot \text{wanted: } P_2$

move 1 $\cdot$ rigid walls: V fixed $\cdot$ sealed: n fixed $\cdot V_1 = V_2$ cancels

Step 2 • Convert to kelvin

$$T_1 = 12.0 + 273.15 = 285.15 \text{ K} \cdot T_2 = 42.0 + 273.15 = 315.15 \text{ K}$$

move 2 $\cdot$ every temperature, before any ratio

 **Sense check:** Volume fixed, temperature changing: Gay-Lussac's case, $P_1/T_1 = P_2/T_2$. The pressure may stay in kPa. Only temperature must change scale. ✓

✓ Guided example: solve and check

$$P_1 / T_1 = P_2 / T_2$$

$$P_1 = 655 \text{ kPa} \cdot T_1 = 285.15 \text{ K} \cdot T_2 = 315.15 \text{ K} \cdot V \text{ fixed} \cdot \text{wanted: } P_2$$

Step 3 • Rearrange for the unknown

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \rightarrow P_2 = P_1 \times \frac{T_2}{T_1}$$

Step 4 • Substitute and check

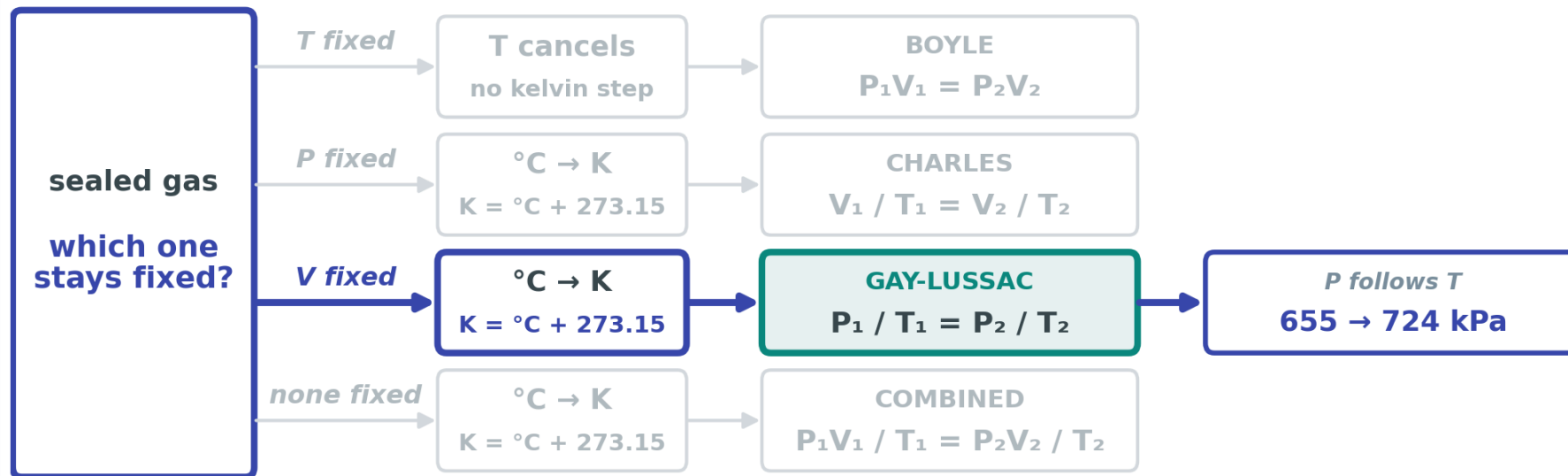
$$P_2 = 655 \text{ kPa} \times \frac{315.15 \text{ K}}{285.15 \text{ K}} = \mathbf{724 \text{ kPa}}$$

 **Sense check:** Move 4: heating a fixed volume raises the pressure, and 724 kPa is above 655 kPa ✓. A Celsius ratio, $655 \times 42.0/12.0 = 2.29 \times 10^3 \text{ kPa}$, would more than triple it.

✓ Guided example: the route on the chart

$$P_2 = 655 \text{ kPa} \times 315.15 \text{ K} / 285.15 \text{ K} = 724 \text{ kPa}$$

V fixed · kelvin step used · found: $P_2 = 724 \text{ kPa}$



🧠 **Sense check:** Rigid walls fix V, so the route runs through the kelvin step to Gay-Lussac's law. ✓

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

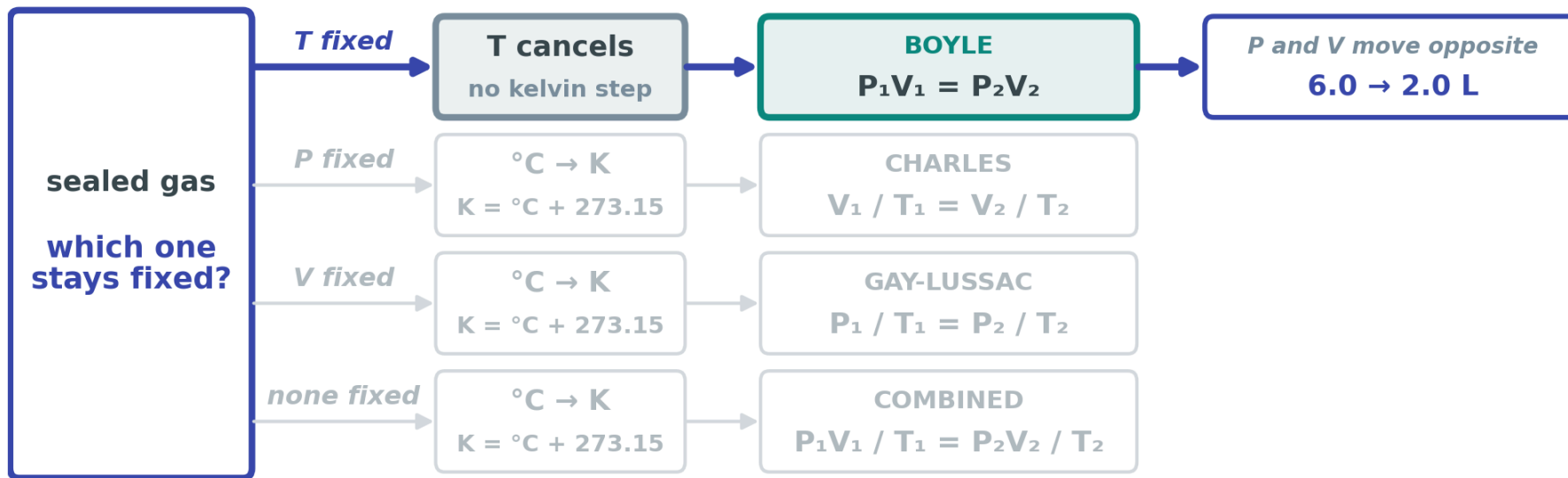
$$V_2 = 6.0 \text{ L} \times \frac{1.0 \text{ atm}}{3.0 \text{ atm}} = 2.0 \text{ L}$$

 **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. ✓

Practice 1: a sealed syringe

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: $V_1 = 36.0 \text{ mL}$ · $P_1 = 745 \text{ mmHg}$ · $V_2 = 60.0 \text{ mL}$ · temperature constant · wanted: P_2

A sealed syringe holds **36.0 mL** of air at **745 mmHg**. At constant temperature, the plunger is pulled back until the air fills **60.0 mL**. What is the new pressure, in mmHg?

- A. 1.24×10^3
- B. 447
- C. 2.68×10^4
- D. 2.90

✓ Practice 1: answer B

$$P_1V_1 = P_2V_2 \rightarrow P_2 = P_1 \times V_1 / V_2$$

temperature constant: T cancels · mmHg and mL may stay as given

$$P_2 = 745 \text{ mmHg} \times \frac{36.0 \text{ mL}}{60.0 \text{ mL}} = \mathbf{447 \text{ mmHg} \rightarrow \text{answer B}}$$

A inverted the volume ratio: $745 \times 60.0/36.0 = 1.24 \times 10^3$, a pressure that rose as the gas expanded. C multiplied P_1 by V_1 and stopped: $745 \times 36.0 = 2.68 \times 10^4$. D swapped P_1 and V_2 when rearranging, $P_2 = V_1V_2/P_1$: $36.0 \times 60.0/745 = 2.90$, in mL^2/mmHg , not a pressure.

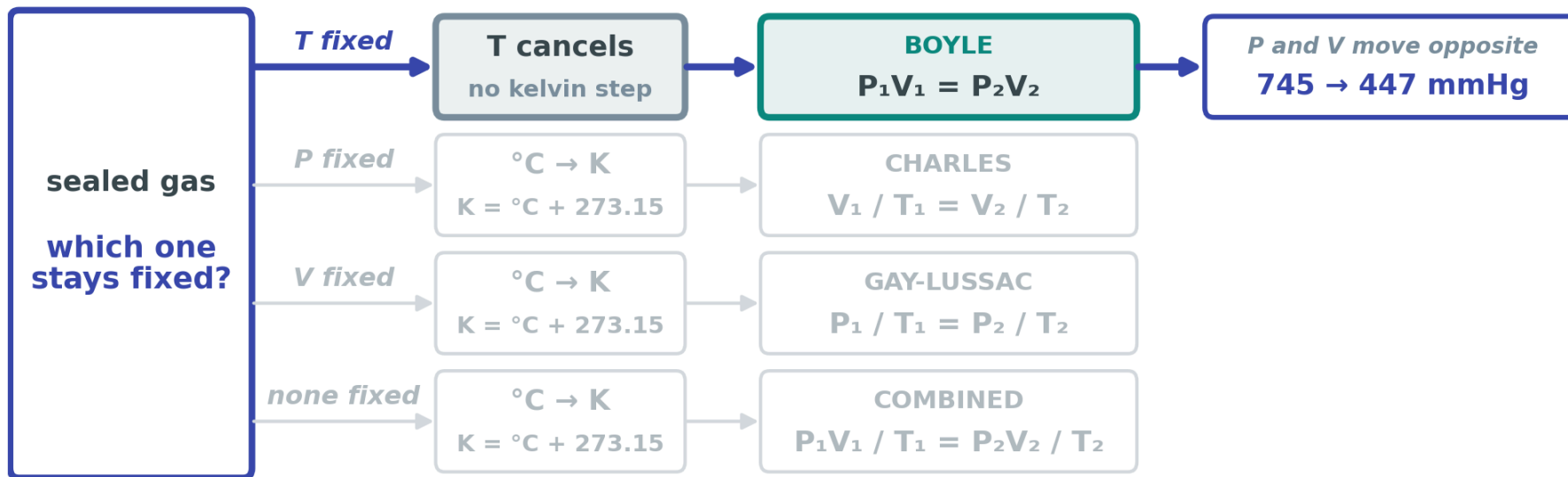
 **Sense check:** The volume grew $60.0/36.0 = 1.67$ times, so the pressure fell by the same factor: $745 \rightarrow 447 \text{ mmHg}$.

✓

✓ Practice 1: the route on the chart

$$P_2 = 745 \text{ mmHg} \times 36.0 \text{ mL} / 60.0 \text{ mL} = 447 \text{ mmHg}$$

T fixed · no kelvin step · found: $P_2 = 447 \text{ mmHg}$ (answer B)



 **Sense check:** Temperature held constant, so the route skips the kelvin step. Pressure and volume stay in the units given. ✓



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

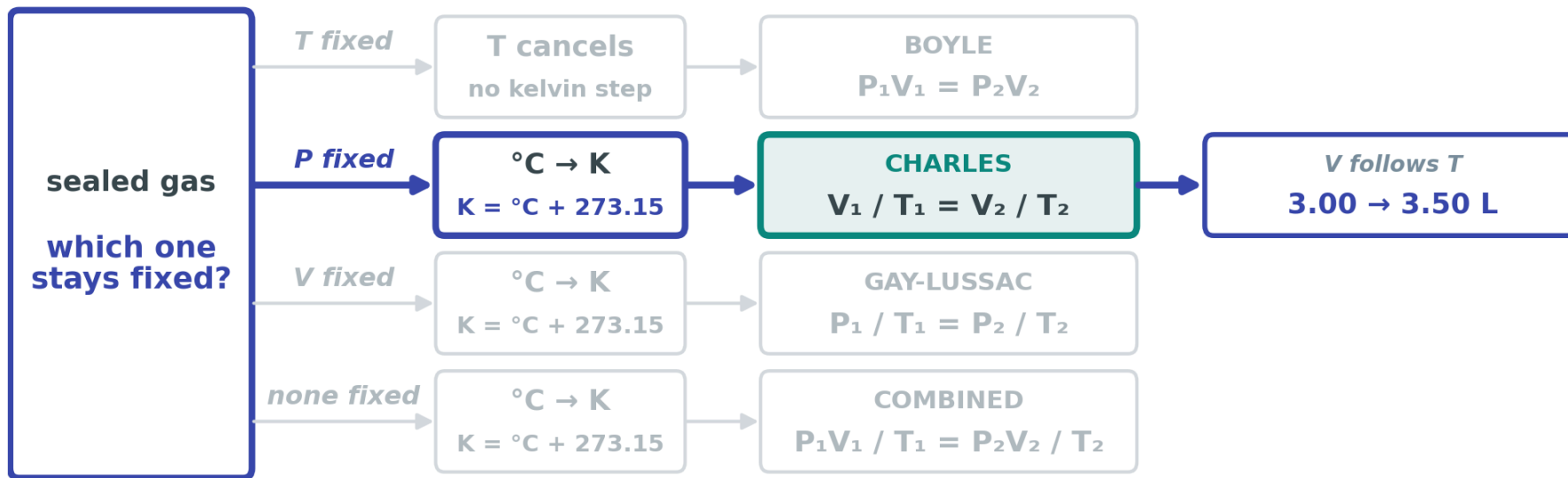
$$V_2 = 3.00 \text{ L} \times \frac{348.15 \text{ K}}{298.15 \text{ K}} = 3.50 \text{ L}$$

 **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 2

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

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 2 answer: C

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

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

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

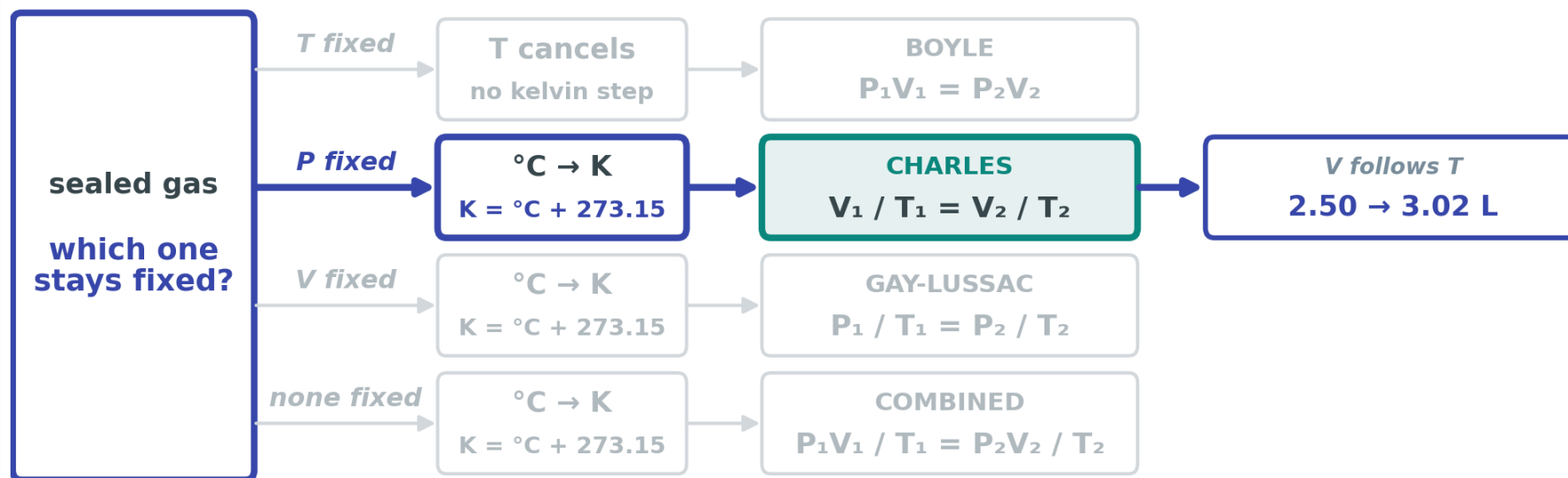
A left the temperatures in Celsius and inverted them: $2.50 \times (17/77) = 0.552 \text{ L}$. B inverted the kelvin ratio: $2.50 \times (290.15/350.15) = 2.07 \text{ L}$, a shrinking gas though it was heated. D left the temperatures in Celsius: $2.50 \times (77/17) = 11.3 \text{ L}$, a ratio that far 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: $2.50 \text{ L} \rightarrow 3.02 \text{ L}$. ✓

✓ Practice 2: the route on the chart

$$V_2 = 2.50 \text{ L} \times 350.15 \text{ K} / 290.15 \text{ K} = 3.02 \text{ L}$$

P fixed · kelvin step used · found: $V_2 = 3.02 \text{ L}$ (answer C)



🧠 **Sense check:** Pressure held constant while the temperature changed: the kelvin step, then Charles's law. ✓



Practice 3: a balloon in the freezer

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: 3.40 L at 23.0 °C → 2.95 L · pressure constant · wanted: T_2 in °C

A party balloon holds **3.40 L** of air at **23.0 °C**. In a freezer, at constant pressure, it shrinks to **2.95 L**. What is the temperature inside the freezer, in °C?

- A. 257
- B. 20.0
- C. 68
- D. -16

✓ Practice 3: answer D

$$V_1 / T_1 = V_2 / T_2 \rightarrow T_2 = T_1 \times V_2 / V_1$$

pressure constant · $T_1 = 23.0 + 273.15 = 296.15 \text{ K}$

$$T_2 = 296.15 \text{ K} \times \frac{2.95 \text{ L}}{3.40 \text{ L}} = 257 \text{ K} \rightarrow 257 - 273.15 = -16 \text{ °C} \rightarrow \text{answer D}$$

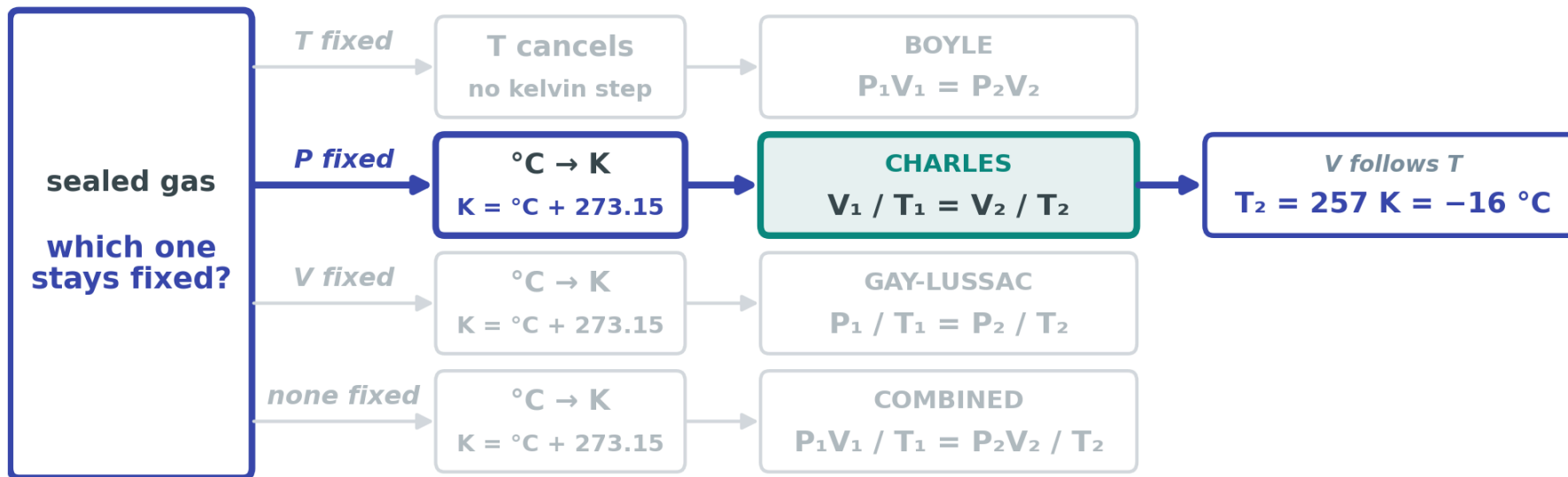
A stopped at kelvin: 257 K is the right temperature on the wrong scale. B put Celsius into the ratio: $23.0 \times 2.95/3.40 = 20.0 \text{ °C}$. C inverted the volume ratio: $296.15 \times 3.40/2.95 = 341 \text{ K} = 68 \text{ °C}$, a gas that warmed as it shrank.

 **Sense check:** The volume fell to 0.868 of its start, so the kelvin temperature did too: below freezing, as a freezer should be. ✓

✓ Practice 3: the route on the chart

$$T_2 = 296.15 \text{ K} \times 2.95 \text{ L} / 3.40 \text{ L} = 257 \text{ K} = -16 \text{ }^{\circ}\text{C}$$

P fixed · kelvin step used, then back to °C · found: $-16 \text{ }^{\circ}\text{C}$ (answer D)



Sense check: Pressure held constant: Charles's law, solved for T_2 . The kelvin answer converts back because the question asked for °C. ✓

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

A common first attempt

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

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

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

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

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

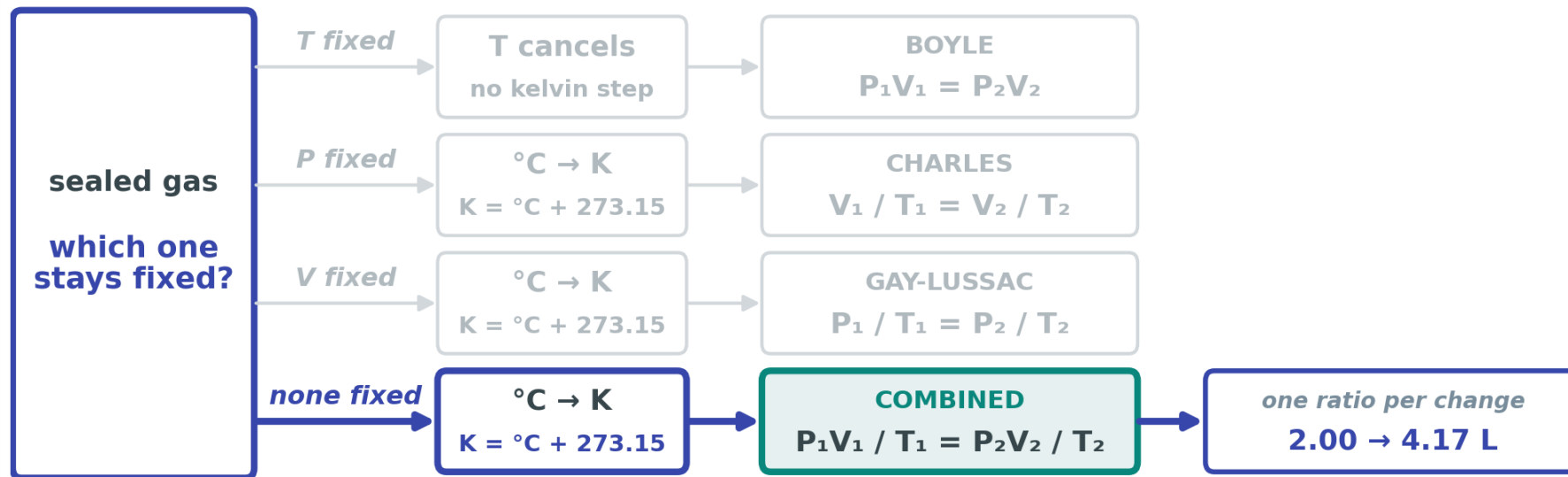
$$V_2 = 2.00 \text{ L} \times \frac{1.00 \text{ atm}}{0.400 \text{ atm}} \times \frac{250.15 \text{ K}}{300.15 \text{ K}} = 4.17 \text{ L}$$

 **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 4: a diver's breath

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: 0.600 L · 2.80 atm · 10.0 °C → 1.50 L · 37.0 °C · wanted: P_2

A diver breathes **0.600 L** of air from a tank at **2.80 atm** and **10.0 °C**. In the lungs, the same air warms to **37.0 °C** and fills **1.50 L**. What is its pressure there, in atm?

- A. 1.23
- B. 4.14
- C. 7.67
- D. 1.02
- E. 1.12

✓ Practice 4: answer A

$$P_2 = P_1 \times (V_1 / V_2) \times (T_2 / T_1)$$

$$\text{nothing fixed} \cdot T_1 = 10.0 + 273.15 = 283.15 \text{ K} \cdot T_2 = 37.0 + 273.15 = 310.15 \text{ K}$$

$$P_2 = 2.80 \text{ atm} \times \frac{0.600 \text{ L}}{1.50 \text{ L}} \times \frac{310.15 \text{ K}}{283.15 \text{ K}} = \mathbf{1.23 \text{ atm} \rightarrow \text{answer A}}$$

B left the temperatures in Celsius: $2.80 \times 0.400 \times (37.0/10.0) = 4.14$. C inverted the volume ratio: $2.80 \times (1.50/0.600) \times (310.15/283.15) = 7.67$. D inverted the temperature ratio: $2.80 \times 0.400 \times (283.15/310.15) = 1.02$. E held temperature constant: $2.80 \times 0.400 = 1.12$.

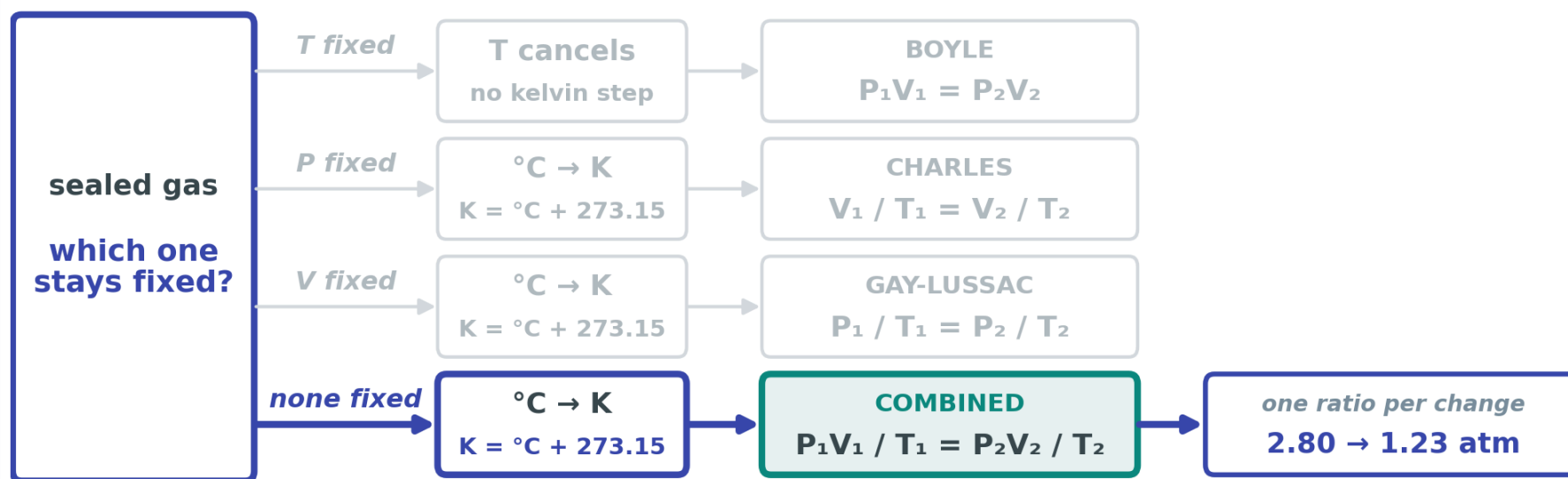
 **Sense check:** Expanding 2.5 times alone would drop the pressure to 1.12 atm. Warming lifts it a little, to 1.23 atm.



✓ Practice 4: the route on the chart

$$P_2 = 2.80 \text{ atm} \times (0.600 \text{ L} / 1.50 \text{ L}) \times (310.15 \text{ K} / 283.15 \text{ K}) = 1.23 \text{ atm}$$

nothing fixed · kelvin step used · found: $P_2 = 1.23 \text{ atm}$ (answer A)



Sense check: Volume and temperature both changed, so the route ends at the combined law with two ratios. ✓

 Practice 5

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

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 5 answer: B

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

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

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

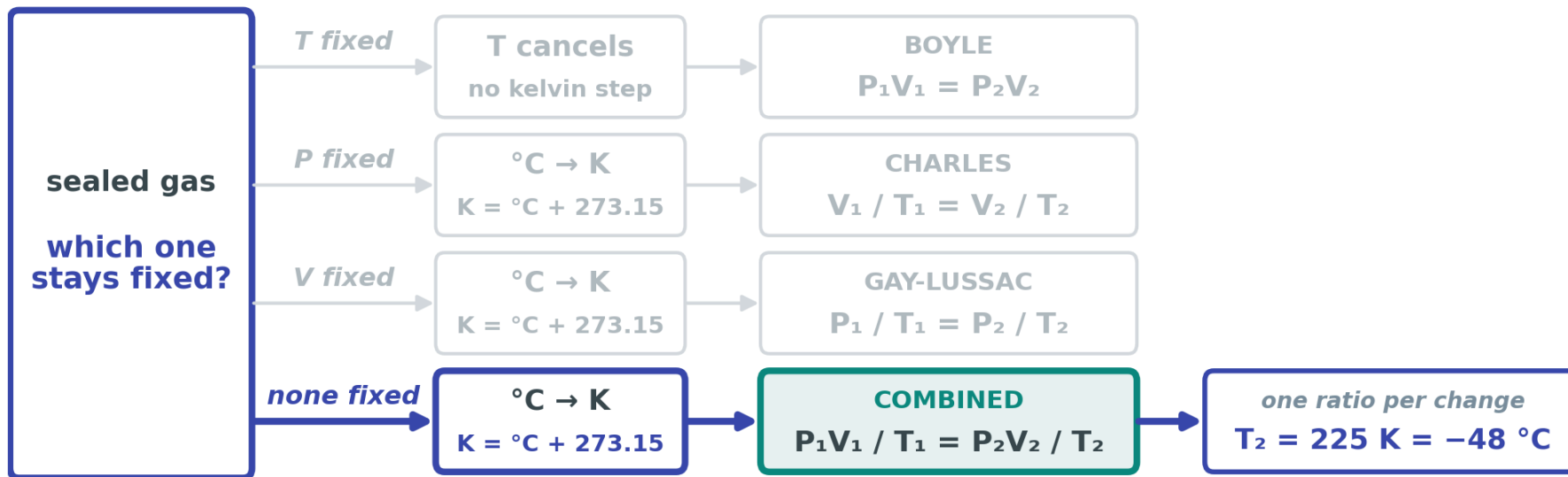
A inverted the PV ratio: $294.15 \times 15.68 / 12.0 = 384 \text{ K} = 111 \text{ }^\circ\text{C}$, a gas that warmed as it rose. C stopped at kelvin: 225 K is the right temperature on the wrong scale. D put Celsius into the ratio: $21.0 \times 12.0 / 15.68 = 16.1 \text{ }^\circ\text{C}$.

 **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. ✓

✓ Practice 5: the route on the chart

$$T_2 = 294.15 \text{ K} \times (0.300 \text{ atm} \times 40.0 \text{ L}) / (0.980 \text{ atm} \times 16.0 \text{ L}) = 225 \text{ K} = -48 \text{ }^{\circ}\text{C}$$

nothing fixed · kelvin step used, then back to $^{\circ}\text{C}$ · found: $-48 \text{ }^{\circ}\text{C}$ (answer B)



 **Sense check:** Pressure and volume both changed, so the combined law solves for T_2 . Kelvin comes out; the question asked for $^{\circ}\text{C}$. ✓

Practice 6: a cold piston warms

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: 1.80 L · 1.45 atm, −20.0 °C → 0.850 atm, 35.0 °C · wanted: V_2

A piston holds **1.80 L** of argon at **1.45 atm and −20.0 °C**. The pressure is lowered to **0.850 atm** while the gas warms to **35.0 °C**. What volume does the argon occupy, in L?

- A. 1.28
- B. 3.07
- C. −5.37
- D. 3.74
- E. 2.52

✓ Practice 6: answer D

$$V_2 = V_1 \times (P_1 / P_2) \times (T_2 / T_1)$$

$$\text{nothing fixed} \cdot T_1 = -20.0 + 273.15 = 253.15 \text{ K} \cdot T_2 = 35.0 + 273.15 = 308.15 \text{ K}$$

$$V_2 = 1.80 \text{ L} \times \frac{1.45 \text{ atm}}{0.850 \text{ atm}} \times \frac{308.15 \text{ K}}{253.15 \text{ K}} = 3.74 \text{ L} \rightarrow \text{answer D}$$

A inverted the pressure ratio: $1.80 \times (0.850/1.45) \times (308.15/253.15) = 1.28$, a gas that shrank as the pressure fell. B held temperature constant: $1.80 \times (1.45/0.850) = 3.07$. C left the temperatures in Celsius: $1.80 \times (1.45/0.850) \times (35.0/-20.0) = -5.37$, a negative volume. E inverted the temperature ratio: $1.80 \times (1.45/0.850) \times (253.15/308.15) = 2.52$.

 **Sense check:** Both changes expand the gas. The pressure fell to 0.586 of its start, and the kelvin temperature rose by a factor of 1.22. So V_2 must beat 3.07 L, the pressure change alone. ✓

Practice 7: a diesel cold start

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: 0.520 L · 0.975 atm · $-15.0\text{ }^{\circ}\text{C}$ → 0.0310 L · $520.\text{ }^{\circ}\text{C}$ · wanted: P_2 in mmHg

On a $-15.0\text{ }^{\circ}\text{C}$ morning, a diesel engine cylinder draws in **0.520 L** of air at **0.975 atm**. The compression stroke squeezes the air to **0.0310 L** and heats it to **$520.\text{ }^{\circ}\text{C}$** . What is its pressure then, in mmHg?

- A. 50.2
- B. 1.24×10^4
- C. -4.31×10^5
- D. 136
- E. 3.82×10^4

✓ Practice 7: answer E

$$P_2 = P_1 \times (V_1 / V_2) \times (T_2 / T_1), \text{ then } 760 \text{ mmHg} = 1 \text{ atm}$$

$$T_1 = -15.0 + 273.15 = 258.15 \text{ K} \cdot T_2 = 520. + 273.15 = 793.15 \text{ K}$$

$$P_2 = 0.975 \text{ atm} \times \frac{0.520 \text{ L}}{0.0310 \text{ L}} \times \frac{793.15 \text{ K}}{258.15 \text{ K}} \times \frac{760 \text{ mmHg}}{1 \text{ atm}}$$

$$= 3.82 \times 10^4 \text{ mmHg} \rightarrow \text{answer E}$$

A stopped in atm: $0.975 \times (0.520/0.0310) \times (793.15/258.15) = 50.2$, the pressure in the wrong unit. B held temperature constant: $0.975 \times (0.520/0.0310) \times 760 = 1.24 \times 10^4$. C left the temperatures in Celsius: $0.975 \times (0.520/0.0310) \times (520./-15.0) \times 760 = -4.31 \times 10^5$. D inverted the volume ratio: $0.975 \times (0.0310/0.520) \times (793.15/258.15) \times 760 = 136$.

 **Sense check:** Squeezing (volume $\div$ 16.8) and heating (kelvin $\times$ 3.07) both raise the pressure, about 50-fold. ✓



Practice 8: from STP to a cold room

$$P_1V_1 / T_1 = P_2V_2 / T_2$$

given: 3.60 L at STP $\rightarrow$ 2.35 atm, $-40.0\text{ }^{\circ}\text{C}$ · wanted: V_2

A sample of argon occupies **3.60 L at STP**. In a **$-40.0\text{ }^{\circ}\text{C}$** cold room, it is compressed to **2.35 atm**. What volume does it occupy there, in L?

- A. 1.31
- B. 7.22
- C. 1.53
- D. 994
- E. 1.79

✓ Practice 8: answer A

$$V_2 = V_1 \times (P_1 / P_2) \times (T_2 / T_1)$$

$$\text{STP: } P_1 = 1 \text{ atm, } T_1 = 273.15 \text{ K} \cdot T_2 = -40.0 + 273.15 = 233.15 \text{ K}$$

$$V_2 = 3.60 \text{ L} \times \frac{1 \text{ atm}}{2.35 \text{ atm}} \times \frac{233.15 \text{ K}}{273.15 \text{ K}} = 1.31 \text{ L} \rightarrow \text{answer A}$$

B inverted the pressure ratio: $3.60 \times (2.35/1) \times (233.15/273.15) = 7.22$, a gas that grew as it was squeezed. C held temperature constant: $3.60 \times (1/2.35) = 1.53$. D took STP pressure as 760 against 2.35 atm: $3.60 \times (760/2.35) \times (233.15/273.15) = 994$. The two pressures must share one unit. E inverted the temperature ratio: $3.60 \times (1/2.35) \times (273.15/233.15) = 1.79$.

 **Sense check:** Compression and cooling both shrink the gas, so V_2 falls below 1.53 L, the pressure change alone. ✓



Check yourself

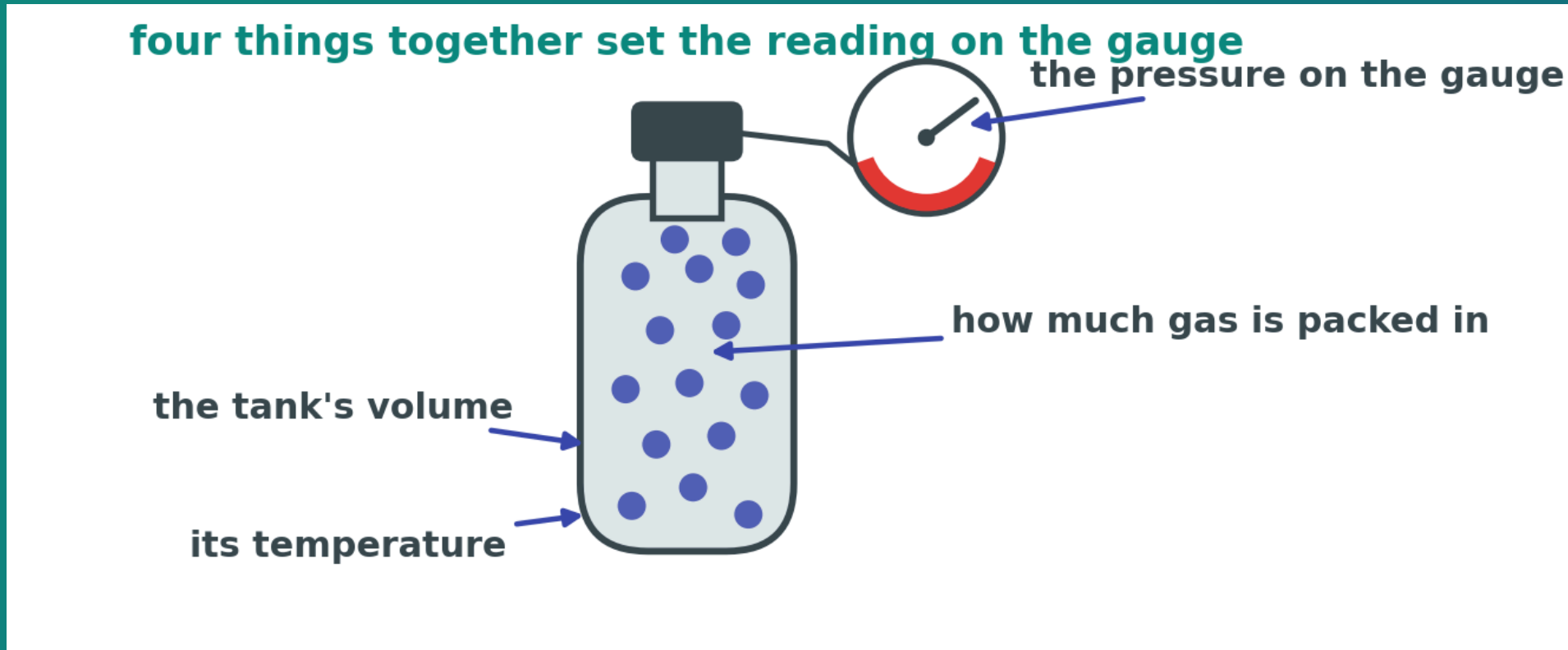
1. A gas is compressed at constant temperature. Which special case is this, and which way does the volume move relative to the pressure?
2. Write $P_1V_1/T_1 = P_2V_2/T_2$ rearranged for V_2 . When a gas is heated, which temperature goes on top of the ratio?

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.

R from one mole at STP

$$R = P V \div (n T)$$

one mole of any gas at STP: 1 atm · 22.414 L · 273.15 K

$$R = \frac{1 \text{ atm} \times 22.414 \text{ L}}{1 \text{ mol} \times 273.15 \text{ K}} = \mathbf{0.08206 \text{ L}\cdot\text{atm/mol}\cdot\text{K}}$$

At STP, one mole of any gas fills 22.4 L. Divide PV by nT for that mole. Every gas gives the same number: the gas constant R.

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

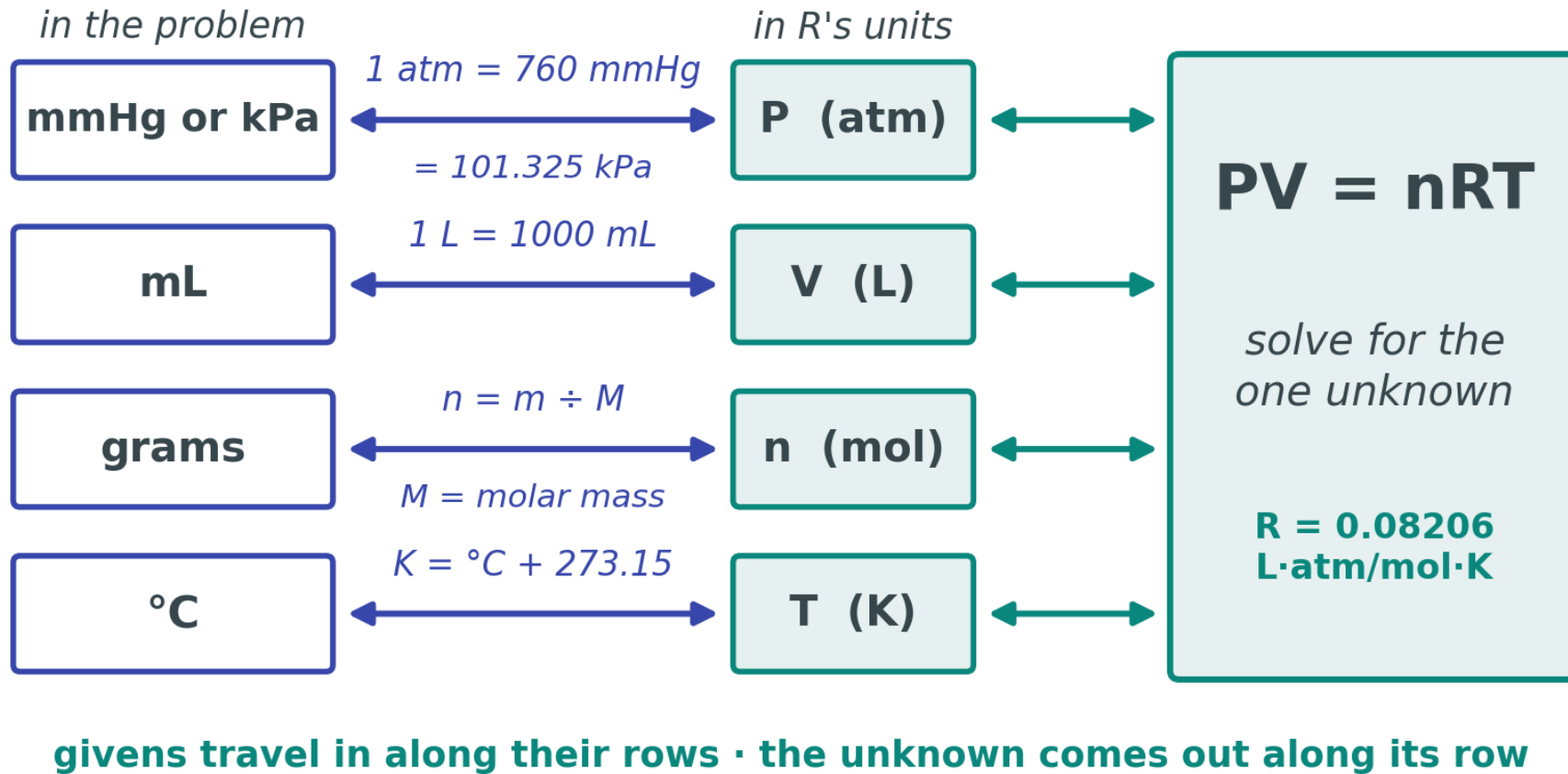
The combined gas law compares a fixed sample before and after a change. The ideal gas law describes a single state and puts the amount of gas, n , directly in.



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.**

The ideal gas law map



Each given travels in along its row and enters in R's units. The one unknown comes out of $PV = nRT$ along its own row. Grams join through $n = m \div M$.



Guided example: a party balloon

$$P V = n R T$$

given: 0.250 mol helium · 1.00 atm · 20.0 °C · wanted: V in L

A party balloon holds **0.250 mol** of helium at **1.00 atm** and **20.0 °C**. What volume does the balloon fill?

Mark which pieces are already in R's units, then solve for V.

✓ Guided example: solution

$$P V = n R T$$

given: 0.250 mol · 1.00 atm · 20.0 °C · wanted: V in L

Step 1 · List the pieces

Step 2 · Rearrange $PV = nRT$

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

Step 3 · Substitute R and cancel units

Step 4 · Check the units and the size

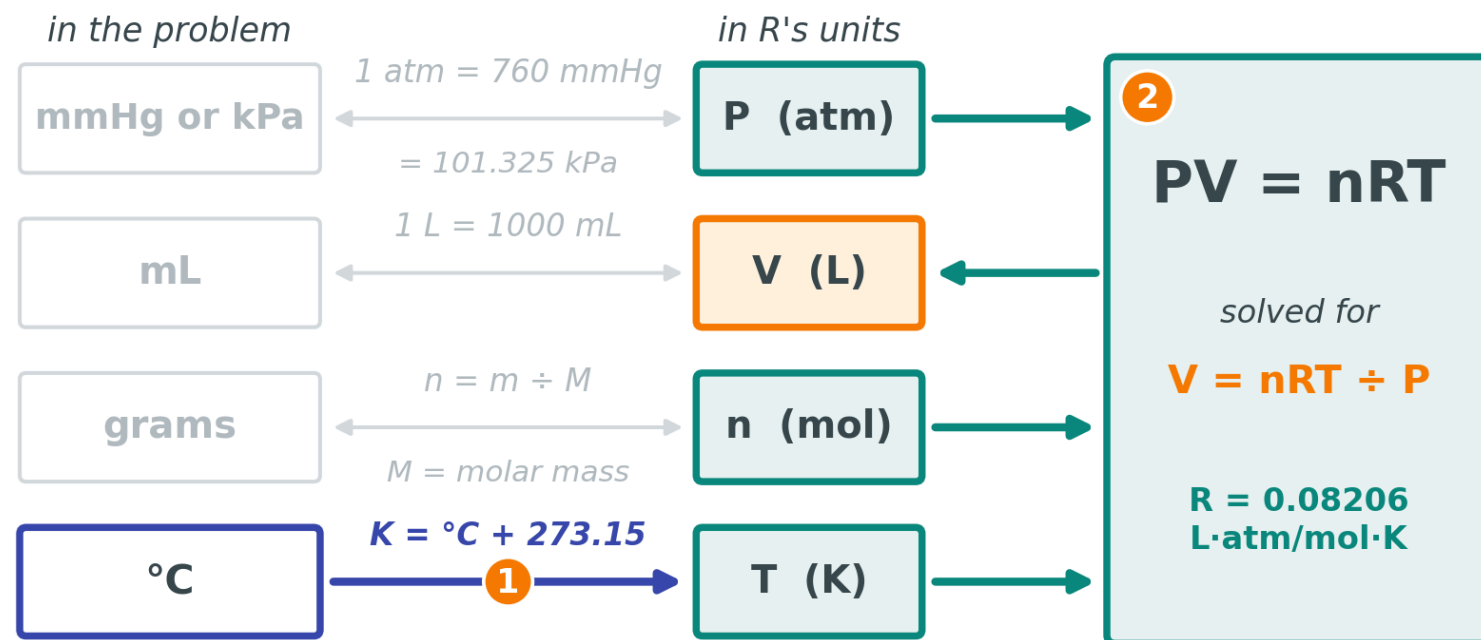
$$V = \frac{0.250 \cancel{\text{mol}} \times 0.08206 \text{ L} \cdot \cancel{\text{atm}} / \cancel{\text{mol}} \cdot \cancel{\text{K}} \times 293.15 \cancel{\text{K}}}{1.00 \cancel{\text{atm}}} = 6.01 \text{ L}$$

🧠 **Sense check:** Only L survives. At STP, $0.250 \times 22.4 = 5.60 \text{ L}$; 20 °C is warmer, so a little more. ✓

✓ Guided example: the route on the map

$$V = nRT \div P$$

given: 0.250 mol · 1.00 atm · 20.0 °C · found: V = 6.01 L



🧠 **Sense check:** Moles and atm were ready. Only °C converted, then V came out of $PV = nRT$. ✓



Practice 1

$$P V = n R T$$

given: 0.0150 mol · 0.350 L · 2.40 atm · wanted: T in K

An aerosol can holds **0.0150 mol** of gas in **0.350 L**. In a fire, at what temperature, in kelvins, does its pressure reach **2.40 atm**?

- A. 409
- B. 1.47×10^{-3}
- C. 682
- D. 10.2

✓ Practice 1 answer: C

$$P V = n R T \rightarrow T = P V \div (n R)$$

given: 2.40 atm · 0.350 L · 0.0150 mol · all three already in R's units · wanted: T in K

$$T = \frac{2.40 \text{ atm} \times 0.350 \text{ L}}{0.0150 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K}} = 682 \text{ K} - \text{answer C}$$

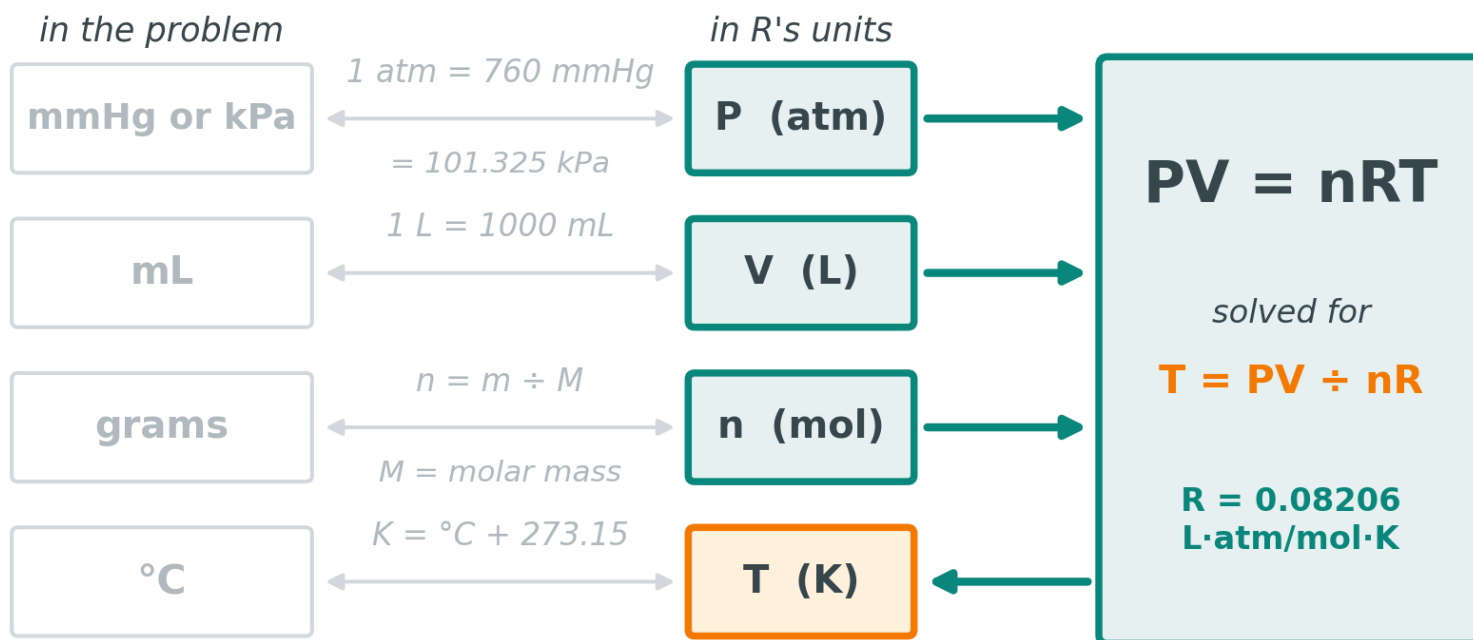
A reported the temperature in Celsius: $682 - 273.15 = 409$. B flipped the rearrangement: $(0.0150 \times 0.08206) \div (2.40 \times 0.350) = 1.47 \times 10^{-3}$. D left n out: $2.40 \times 0.350 \div 0.08206 = 10.2$.

🧠 **Sense check:** 682 K is about 409 °C, the heat of a fire. Only K survives the cancel. ✓

✓ Practice 1: the route on the map

$$T = PV \div nR$$

given: $2.40 \text{ atm} \cdot 0.350 \text{ L} \cdot 0.0150 \text{ mol} \cdot \text{found: } T = 682 \text{ K}$



🧠 **Sense check:** All three givens were already in R's units. One move: T came straight out of $PV = nRT$. ✓



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

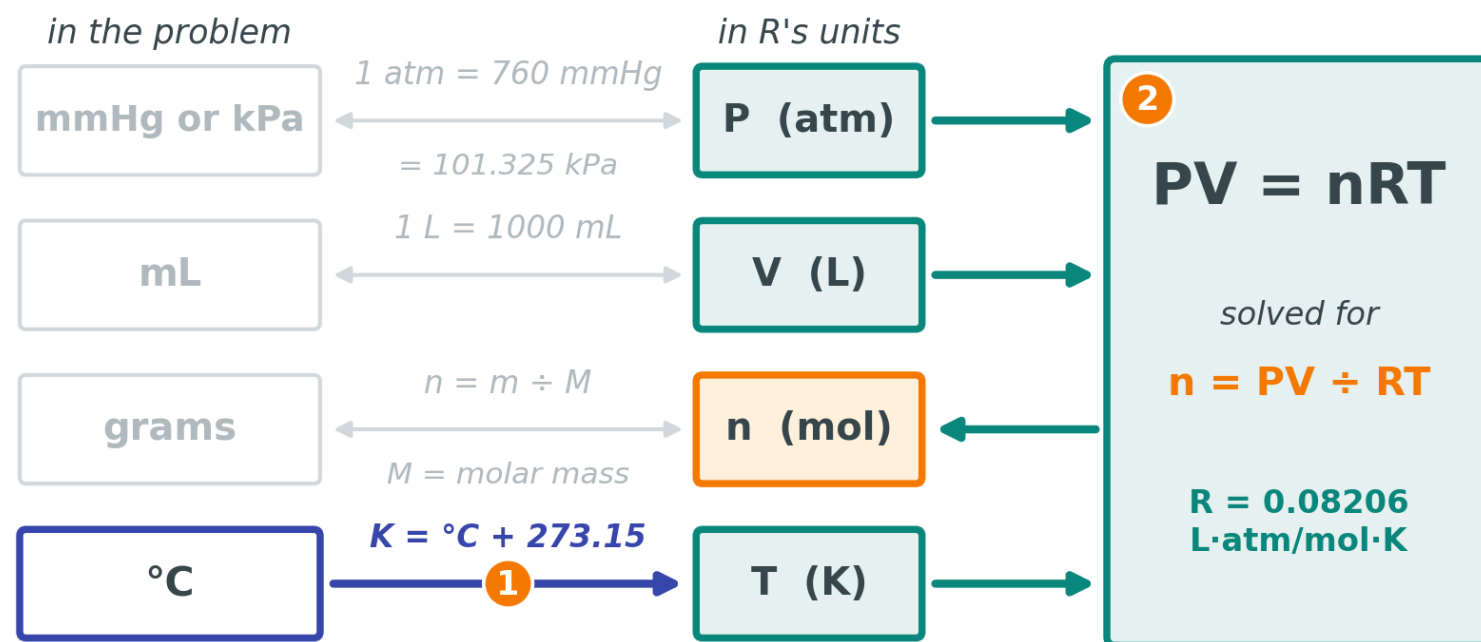
$$n = \frac{2.00 \text{ atm} \times 5.00 \text{ L}}{0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} \times 298.15 \text{ K}} = 0.409 \text{ mol}$$

 **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

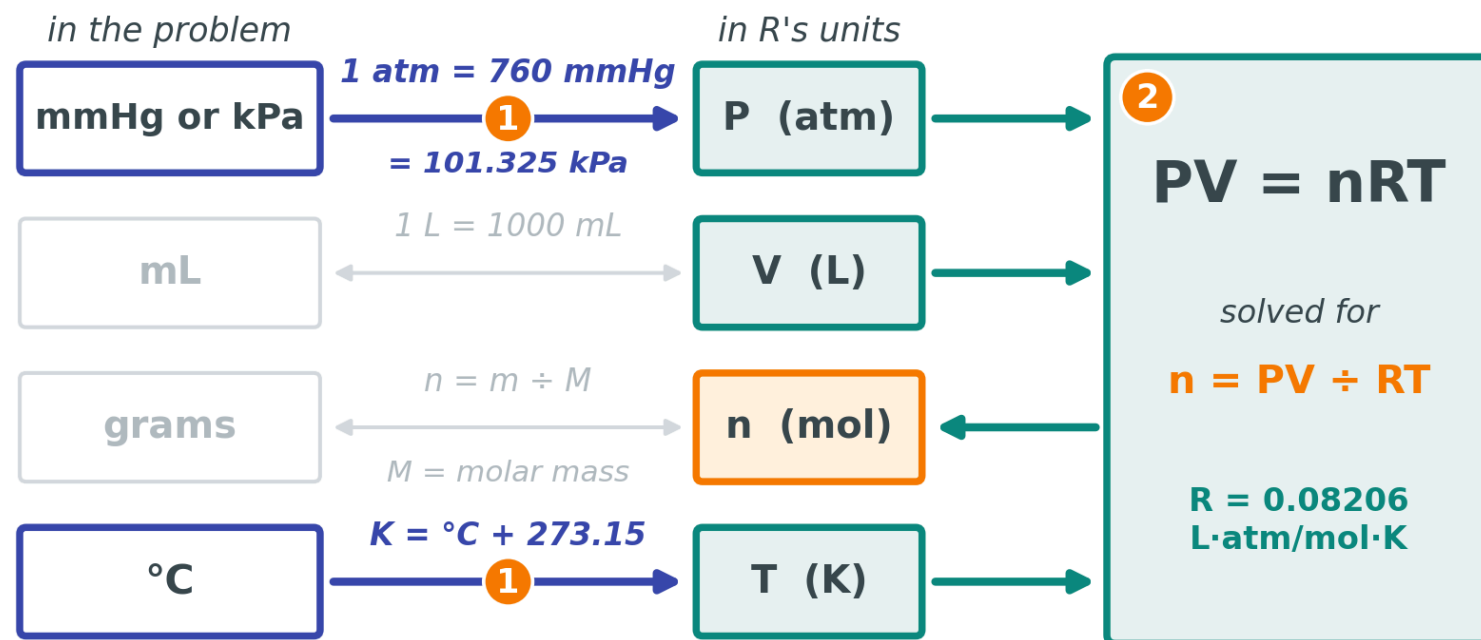
$$n = \frac{1.50 \text{ atm} \times 8.00 \text{ L}}{0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} \times 300.15 \text{ K}} = 0.487 \text{ mol}$$

 **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 2

$$P V = n R T$$

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 2 answer: B

$$P V = n R T$$

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

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

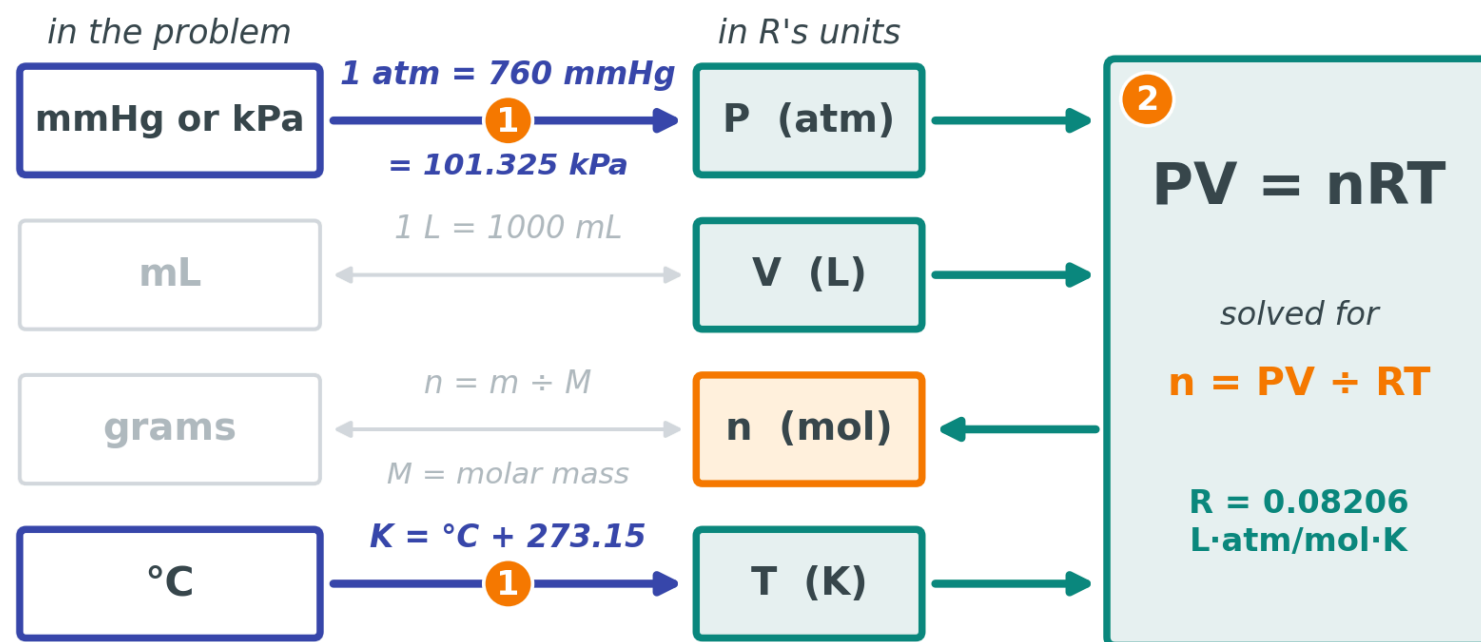
A flipped the rearrangement: $(0.08206 \times 308.15) \div (1.25 \times 4.00) = 5.06$. C left the temperature in Celsius: $1.25 \times 4.00 \div (0.08206 \times 35.0) = 1.74 \text{ mol}$. D left the pressure in mmHg: $950. \times 4.00 \div (0.08206 \times 308.15) = 150. \text{ mol}$.

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

✓ Practice 2: the route on the map

$$n = PV \div RT$$

given: 950. mmHg · 4.00 L · 35.0 °C · found: $n = 0.198$ mol



🧠 **Sense check:** mmHg and °C converted first; the liters were ready. Then n came out of $PV = nRT$. ✓

Molar mass and density

from $PV = nRT$, with $n = m \div M \cdot M = m R T \div (P V)$
moles = mass $\div$ molar mass $\cdot$ 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

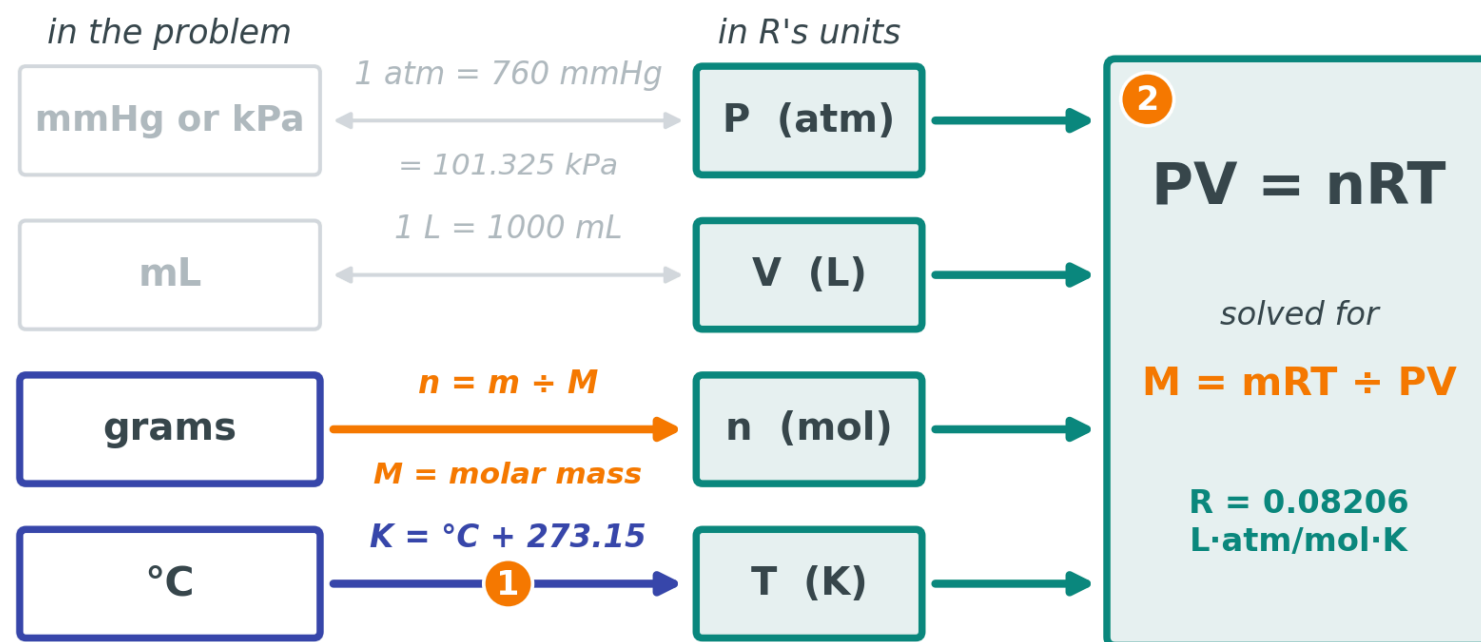
Step 4 · Check the units and the size

 **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 3

$$P V = n R T$$

given: 4.50 L · 3.20 atm · 16.0 °C · Ne 20.18 g/mol · wanted: mass in g

A **4.50 L** cylinder of neon for sign tubes reads **3.20 atm** at **16.0 °C**. What mass of neon, in grams, does it hold?

- A. 12.2
- B. 0.607
- C. 0.0301
- D. 221

✓ Practice 3 answer: A

$$n = P V \div (R T), \text{ then mass} = n \times M$$

two moves · 16.0 + 273.15 = 289.15 K · Ne 20.18 g/mol · wanted: mass in g

$$n = \frac{3.20 \text{ atm} \times 4.50 \text{ L}}{0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} \times 289.15 \text{ K}} = 0.607 \text{ mol Ne}$$

$$0.607 \text{ mol Ne} \times \frac{20.18 \text{ g Ne}}{1 \text{ mol Ne}} = 12.2 \text{ g} - \text{answer A}$$

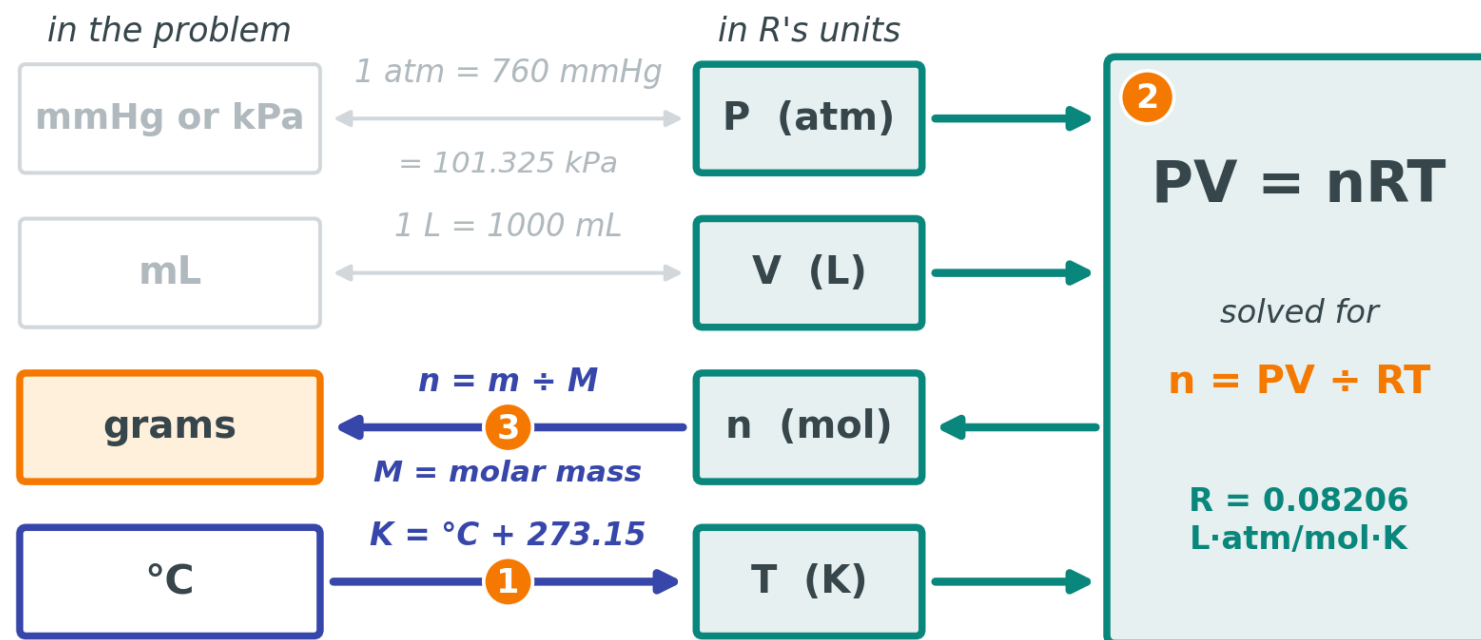
B stopped at moles: 0.607. C divided by the molar mass: $0.607 \div 20.18 = 0.0301$. D left the temperature in Celsius: $3.20 \times 4.50 \div (0.08206 \times 16.0) = 10.97 \text{ mol}$, $\times 20.18 = 221$.

 **Sense check:** About 0.6 mol of a 20 g/mol gas is about 12 g. ✓

✓ Practice 3: the route on the map

$$n = PV \div RT, \text{ then mass} = n \times M$$

given: 4.50 L · 3.20 atm · 16.0 °C · found: 0.607 mol → 12.2 g Ne



🧠 **Sense check:** n came out of $PV = nRT$, then kept going along its row to grams. ✓

 Practice 4

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

given: $d = 1.97 \text{ g/L} \cdot 625 \text{ mmHg} \cdot 22.0 \text{ }^\circ\text{C}$ · 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 °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 4 answer: D

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

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

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

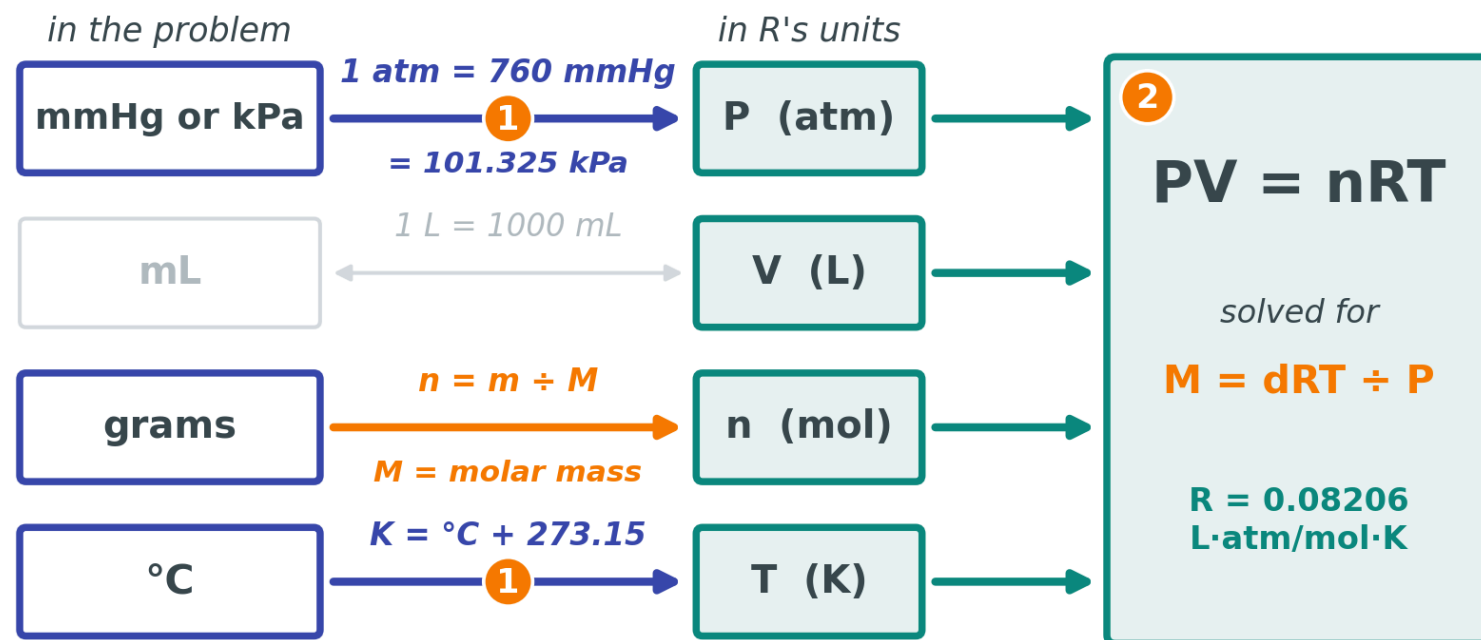
A stopped before dividing by P: $1.97 \times 0.08206 \times 295.15 = 47.7$, as if the pressure were 1 atm. B left the temperature in Celsius: $1.97 \times 0.08206 \times 22.0 \div 0.822 = 4.32$. C flipped the mmHg factor: $760/625 = 1.216 \text{ atm}$, and $47.7 \div 1.216 = 39.2$.

🧠 **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. ✓

✓ Practice 4: the route on the map

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

given: 1.97 g/L · 625 mmHg · 22.0 °C · found: M = 58.0 g/mol



🧠 **Sense check:** Density carries grams and liters together. mmHg and °C converted; M came out. ✓



Practice 5

$$P V = n R T$$

given: 8.50 g NH_3 · 97.5 kPa · 24.0 °C · NH_3 17.03 g/mol · wanted: V in L

A refrigeration line leaks **8.50 g** of ammonia into a sealed collection bag at **97.5 kPa** and **24.0 °C**. What volume does the gas fill, in liters?

- A. 215
- B. 12.6
- C. 0.125
- D. 11.2
- E. 1.02

✓ Practice 5 answer: B

$$V = n R T \div P$$

$$8.50 \text{ g} \div 17.03 \text{ g/mol} = 0.499 \text{ mol} \cdot 97.5 \text{ kPa} \div 101.325 = 0.962 \text{ atm} \cdot 24.0 + 273.15 = 297.15 \text{ K}$$

$$V = \frac{0.499 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 297.15 \text{ K}}{0.962 \text{ atm}} = 12.6 \text{ L} - \text{answer B}$$

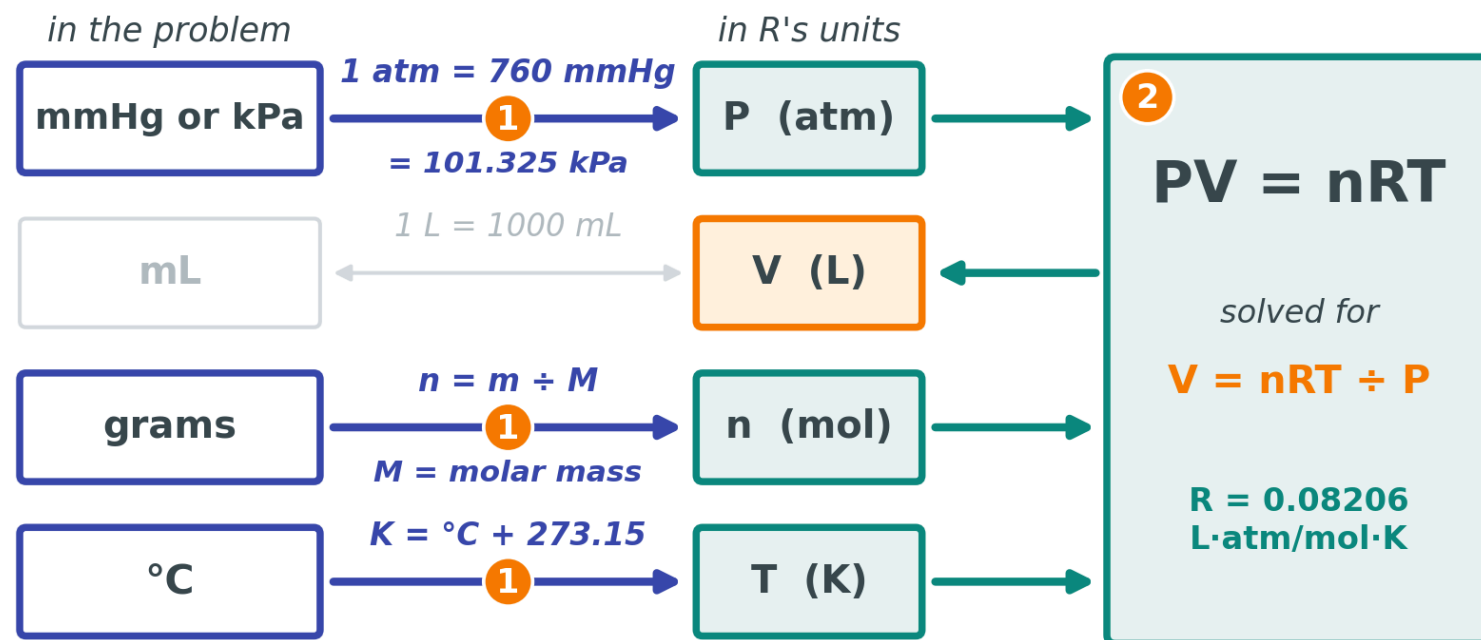
A skipped the molar mass: $8.50 \times 0.08206 \times 297.15 \div 0.962 = 215$. C left the pressure in kPa: $0.499 \times 0.08206 \times 297.15 \div 97.5 = 0.125$. D used 22.4 L/mol, which holds only at STP: $0.499 \times 22.4 = 11.2$. E left the temperature in Celsius: $0.499 \times 0.08206 \times 24.0 \div 0.962 = 1.02$.

🧠 **Sense check:** Half a mole fills more than half of 22.4 L here: the gas is warmer than 0 °C and below 1 atm. ✓

✓ Practice 5: the route on the map

$$V = nRT \div P$$

given: 8.50 g NH₃ · 97.5 kPa · 24.0 °C · found: V = 12.6 L



🧠 **Sense check:** Three conversions came first: grams, kPa and °C. Then V came out of $PV = nRT$. ✓

 Practice 6

$$P V = n R T$$

given: 9.60 g O₂ · 2.50 L · 15.0 °C · wanted: P in atm

A rigid **2.50 L** steel bottle holds **9.60 g** of oxygen gas at **15.0 °C**. What pressure does the oxygen exert, in atm?

- A. 90.8
- B. 0.148
- C. 2.84
- D. 5.67

✓ Practice 6 answer: C

$$n = m \div M, \text{ then } P = n R T \div V$$

$$\text{O}_2 = 2 \times 16.00 = 32.00 \text{ g/mol} \cdot 9.60 \text{ g} \div 32.00 \text{ g/mol} = 0.300 \text{ mol O}_2 \cdot 15.0 + 273.15 = 288.15 \text{ K}$$

$$P = \frac{0.300 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 288.15 \text{ K}}{2.50 \text{ L}} = 2.84 \text{ atm} \rightarrow \text{answer C}$$

A skipped the molar mass: $9.60 \times 0.08206 \times 288.15 \div 2.50 = 90.8$. B left the temperature in Celsius: $0.300 \times 0.08206 \times 15.0 \div 2.50 = 0.148$. D used 16.00 g/mol, one O atom: $0.600 \times 0.08206 \times 288.15 \div 2.50 = 5.67$.

 **Sense check:** Oxygen gas is O₂, 32.00 g per mole. Three tenths of a mole in 2.5 L near room temperature gives a few atm. ✓



Practice 7

$$P V = n R T$$

given: 180. mL · 0.00620 mol · 684 torr · wanted: T in °C

A sealed **180. mL** bulb holds **0.00620 mol** of argon. Set in a heating block, it reads **684 torr**. What is the temperature of the block, in °C?

- A. 45.3
- B. 318
- C. 120.
- D. 592
- E. 3.18×10^5

✓ Practice 7 answer: A

$$T = P V \div (n R), \text{ then } T(^{\circ}\text{C}) = T(\text{K}) - 273.15$$

$$684 \text{ torr} \times (1 \text{ atm} / 760 \text{ torr}) = 0.900 \text{ atm} \cdot 180. \text{ mL} = 0.180 \text{ L} \cdot 0.00620 \text{ mol}$$

$$T = \frac{0.900 \text{ atm} \times 0.180 \text{ L}}{0.00620 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K}} = 318.4 \text{ K}$$

$$318.4 \text{ K} - 273.15 = \mathbf{45.3^{\circ}\text{C} \rightarrow \text{answer A}}$$

B stopped in kelvins: 318. C flipped the torr factor: $760/684 = 1.111$ atm gives 393 K, or 120. °C. D added 273.15 instead of subtracting: $318.4 + 273.15 = 592$. E left the volume in mL: $0.900 \times 180. \div (0.00620 \times 0.08206) - 273.15 = 3.18 \times 10^5$.

 **Sense check:** At STP, 0.00620 mol fills $0.00620 \times 22.4 = 0.139$ L at 1 atm. Here $PV = 0.900 \times 0.180 = 0.162$, more than 0.139, so T sits above 273 K by that ratio: $0.162 \div 0.139 \times 273 \text{ K} \approx 318 \text{ K}$. ✓



Practice 8

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

given: CH_4 16.04 g/mol · 104.5 kPa · 30.0 °C · wanted: d in g/L

Natural gas, nearly pure methane, flows through a city main at **104.5 kPa** and **30.0 °C**. What is its density there, in g/L?

- A. 1.50
- B. 0.665
- C. 67.4
- D. 0.716
- E. 6.72

✓ Practice 8 answer: B

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

$$104.5 \text{ kPa} \div 101.325 \text{ kPa/atm} = 1.031 \text{ atm} \cdot 30.0 + 273.15 = 303.15 \text{ K} \cdot \text{CH}_4 16.04 \text{ g/mol}$$

$$d = \frac{1.031 \text{ atm} \times 16.04 \text{ g/mol}}{0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K} \times 303.15 \text{ K}} = 0.665 \text{ g/L} \rightarrow \text{answer B}$$

A inverted the formula, $RT/(PM)$: $24.88 \div 16.54 = 1.50 \text{ L/g}$. C left the pressure in kPa: $104.5 \times 16.04 \div (0.08206 \times 303.15) = 67.4$. D used the STP molar volume: $16.04 \div 22.4 = 0.716$, true only at 0°C and 1 atm . E left the temperature in Celsius: $1.031 \times 16.04 \div (0.08206 \times 30.0) = 6.72$.

 **Sense check:** The pressure is 1.031 times 1 atm, but the kelvin temperature is 1.11 times 273.15 K. Warmth wins, so the density falls just below the STP value of 0.716 g/L. ✓

Check yourself

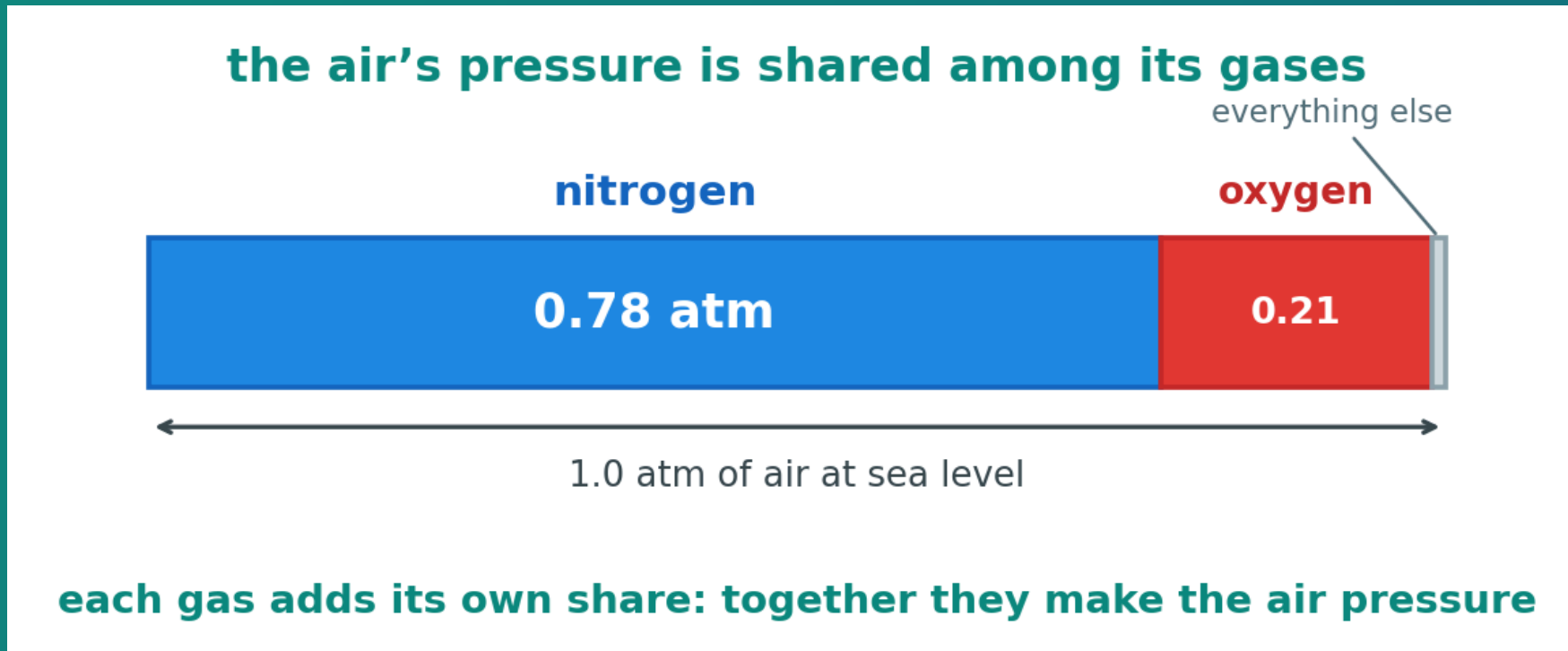
1. A gas occupies 6.00 L at 1.20 atm and 300. K. Rearrange $PV = nRT$ for n , then state which units of R cancel and which unit survives.
2. A pressure reads 745 mmHg and a temperature reads 18 °C. Convert each to the units R requires before either enters the equation.

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.

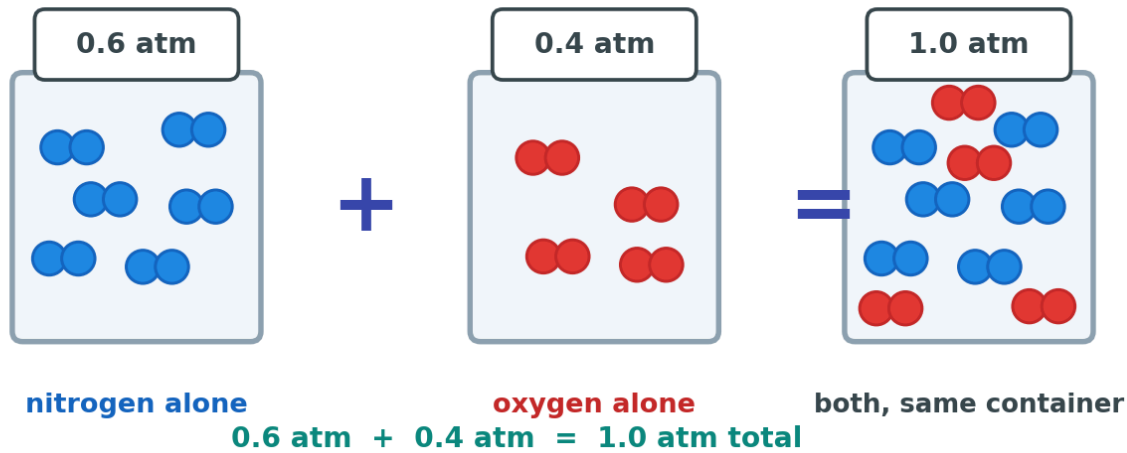
The air you breathe



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



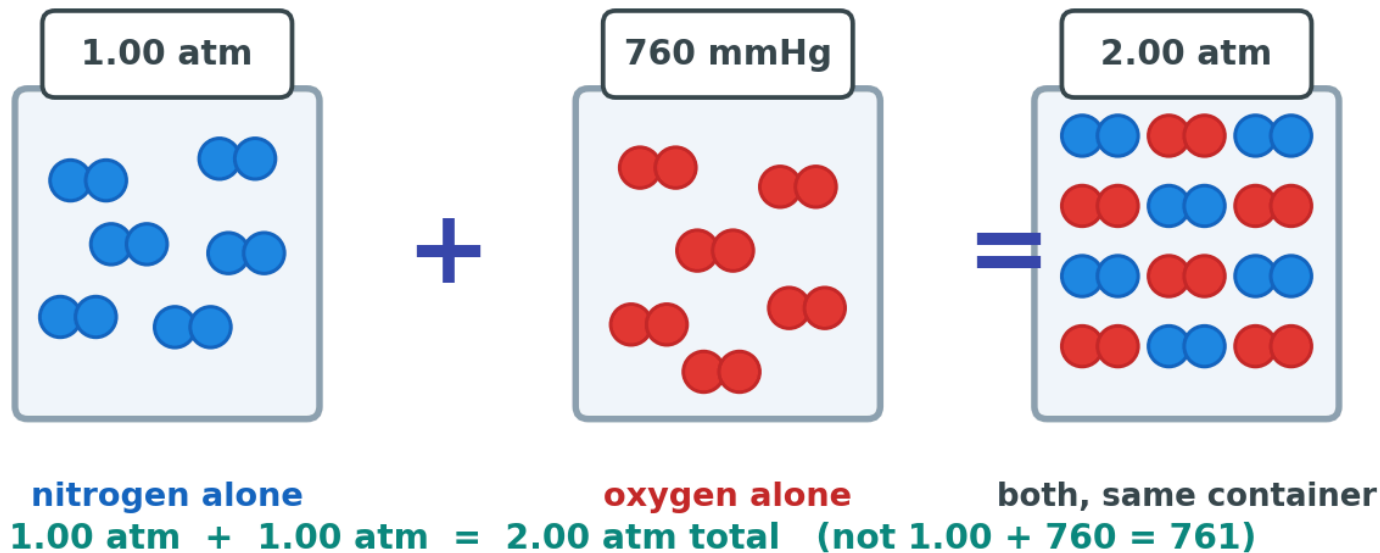
Put two gases into one container and neither disturbs the other. Each keeps the pressure it had alone. The total is the sum of these partial pressures: Dalton's law.

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

Pressures add in one unit

760 mmHg is 1.00 atm: convert to one unit, then add



Dalton's law adds pressure readings, so every reading needs the same unit. Convert first with $760 \text{ mmHg} = 1 \text{ atm}$. The oxygen's 760 mmHg is 1.00 atm, and the total is 2.00 atm.

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

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

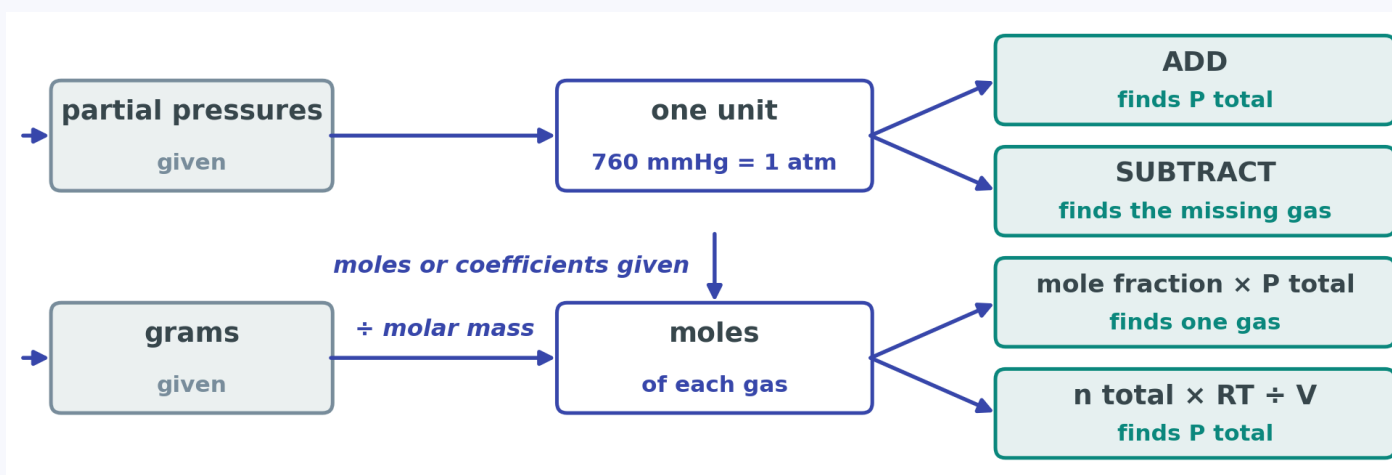
$$\text{mole fraction} = \text{moles of that gas} \div \text{total moles of gas}$$

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, take a mole fraction, or use $n_{\text{total}}RT \div V$.
4. **Check:** partials rebuild the total.





Guided example: a cylinder of synthetic air

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

given: N_2 3268 mmHg · O_2 1.20 atm · wanted: P_{total} in atm

A cylinder of synthetic air for a hospital lab holds nitrogen and oxygen. Nitrogen contributes **3268 mmHg** and oxygen **1.20 atm**. Find the total pressure in the cylinder, in atm.

The two readings carry different units. List the gases and write Dalton's law first.

✓ Guided example: solution

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

given: N₂ 3268 mmHg · O₂ 1.20 atm · wanted: P_{total} in atm

Step 1 · List every gas

Step 2 · Write Dalton's law

Step 3 · Find the unknown

Two moves are needed. Move 1: put nitrogen in atm.

$$P_{\text{N}_2} = 3268 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 4.30 \text{ atm}$$

$$P_{\text{total}} = 4.30 \text{ atm} + 1.20 \text{ atm} = \mathbf{5.50 \text{ atm}}$$

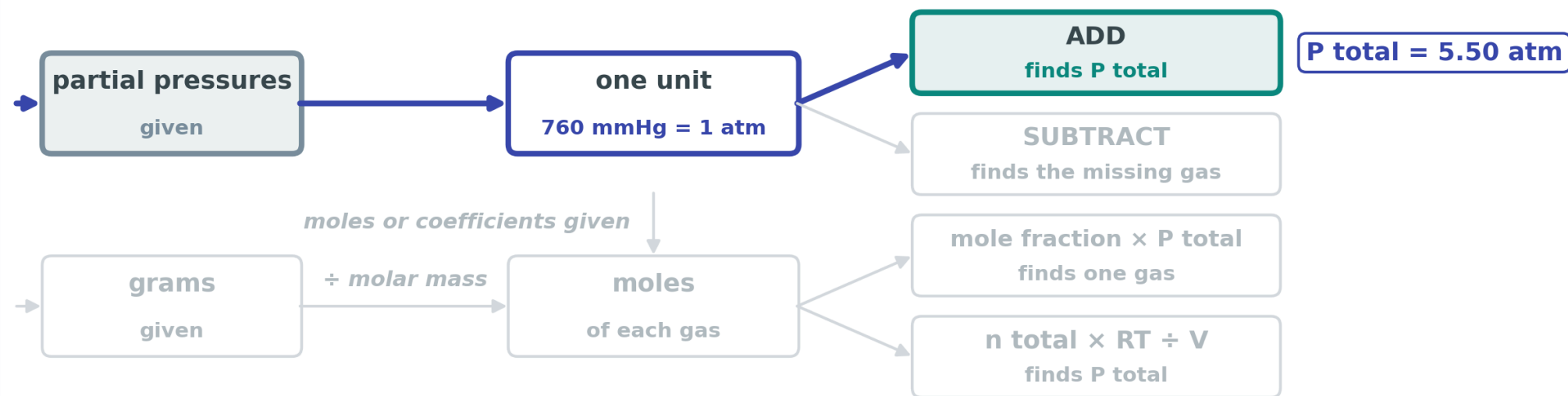
Step 4 · Check

 **Sense check:** Redo it in mmHg: 1.20 atm × 760 = 912 mmHg, and 3268 + 912 = 4180 mmHg, which is 5.50 atm. Adding the bare numbers 3268 and 1.20 mixes two units and gives no pressure at all. ✓

✓ Guided example: the route on the map

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

given: N_2 3268 mmHg · O_2 1.20 atm · found: $P_{\text{total}} = 5.50$ atm



 **Sense check:** Partial pressures were given, so the route stays in the top lane: one unit, then add. Grams and moles never came up. ✓



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

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

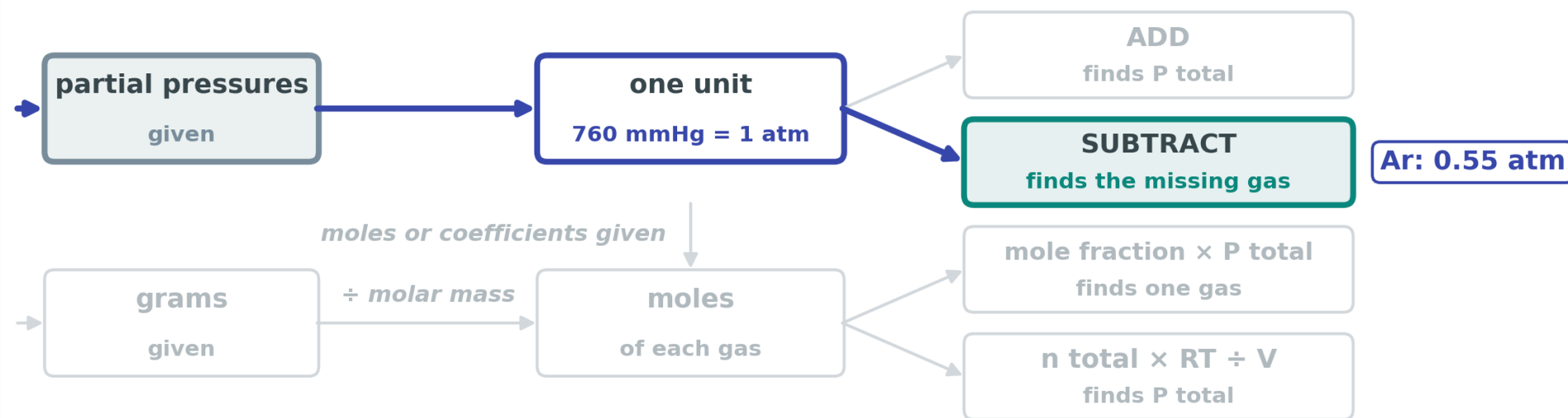
Step 4 · Check

🧠 **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

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

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

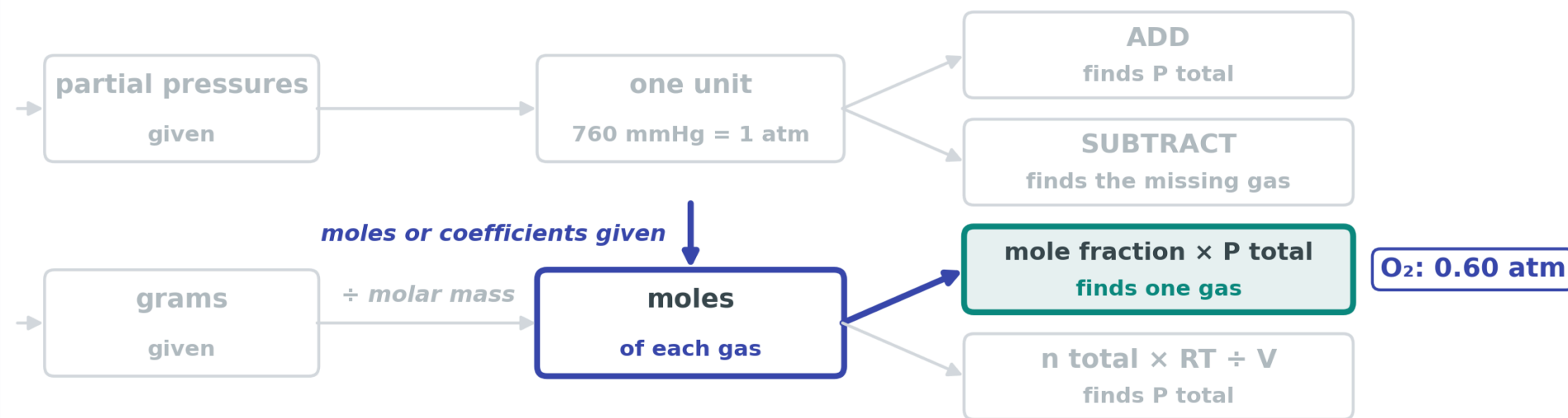
Step 4 · Check

 **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

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

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

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

The route on the map



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



Practice 2

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

given: $P_{\text{total}} = 6.50 \text{ atm}$ · O_2 1330 mmHg · wanted: $P(\text{He})$

For a deep dive, a scuba diver breathes a mixture of helium and oxygen at a total pressure of **6.50 atm**. The oxygen's partial pressure is **1330 mmHg**. What is the partial pressure of helium, in mmHg?

- A. 2470
- B. 3610
- C. 6270
- D. 4940

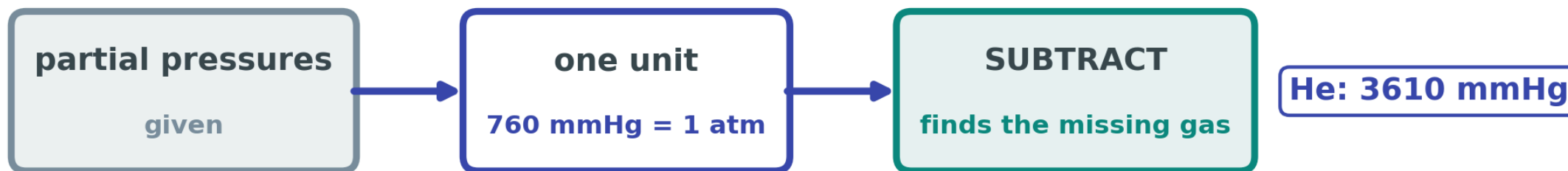
✓ Practice 2 answer: B

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

given: $P_{\text{total}} = 6.50 \text{ atm}$ · O_2 1330 mmHg · wanted: $P(\text{He})$

$$6.50 \text{ atm} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} = 4940 \text{ mmHg} \cdot P_{\text{He}} = 4940 - 1330 = \mathbf{3610 \text{ mmHg}} \rightarrow \text{answer B}$$

The route on the map



Sense check: Add the partials back: $3610 + 1330 = 4940 \text{ mmHg}$, which is 6.50 atm. ✓



Practice 3

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

given: 1.20 mol N₂ · 0.45 mol O₂ · 0.15 mol Ar · P_{total} = 720 mmHg · wanted: P(O₂)

A sealed calibration cylinder holds **1.20 mol of nitrogen**, **0.45 mol of oxygen**, and **0.15 mol of argon** at a total pressure of **720 mmHg**. What is the partial pressure of oxygen, in mmHg?

- A. 240
- B. 480
- C. 270
- D. 180

✓ Practice 3 answer: D

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

$$1.20 + 0.45 + 0.15 = 1.80 \text{ mol total} \cdot P_{\text{total}} = 720 \text{ mmHg}$$

$$X_{\text{O}_2} = \frac{0.45 \text{ mol O}_2}{1.80 \text{ mol total}} = 0.25 \cdot P_{\text{O}_2} = 0.25 \times 720 \text{ mmHg} = \mathbf{180 \text{ mmHg} \rightarrow \text{answer D}}$$

The route on the map



 **Sense check:** Oxygen is one-fourth of the molecules, so it carries one-fourth of 720 mmHg. ✓



Worked example 3: a reaction sets the shares



given: the product gases reach $P_{\text{total}} = 867.0 \text{ mmHg}$ · wanted: $P(\text{NH}_3)$ and $P(\text{H}_2\text{S})$

Ammonium sulfide decomposes completely inside a sealed vessel. The product gases reach a total pressure of **867.0 mmHg**. Find the partial pressure of each gas.

Read the amounts from the coefficients.

✓ Worked example 3: solution



given: the product gases reach $P_{\text{total}} = 867.0 \text{ mmHg}$ · wanted: $P(\text{NH}_3)$ and $P(\text{H}_2\text{S})$

Step 1 · List every gas

Step 2 · Write Dalton's law

Step 3 · Find the unknown

$$P_{\text{NH}_3} = \frac{2 \text{ mol NH}_3}{3 \text{ mol gas}} \times 867.0 \text{ mmHg} = \mathbf{578.0 \text{ mmHg}}$$

$$P_{\text{H}_2\text{S}} = \frac{1 \text{ mol H}_2\text{S}}{3 \text{ mol gas}} \times 867.0 \text{ mmHg} = \mathbf{289.0 \text{ mmHg}}$$

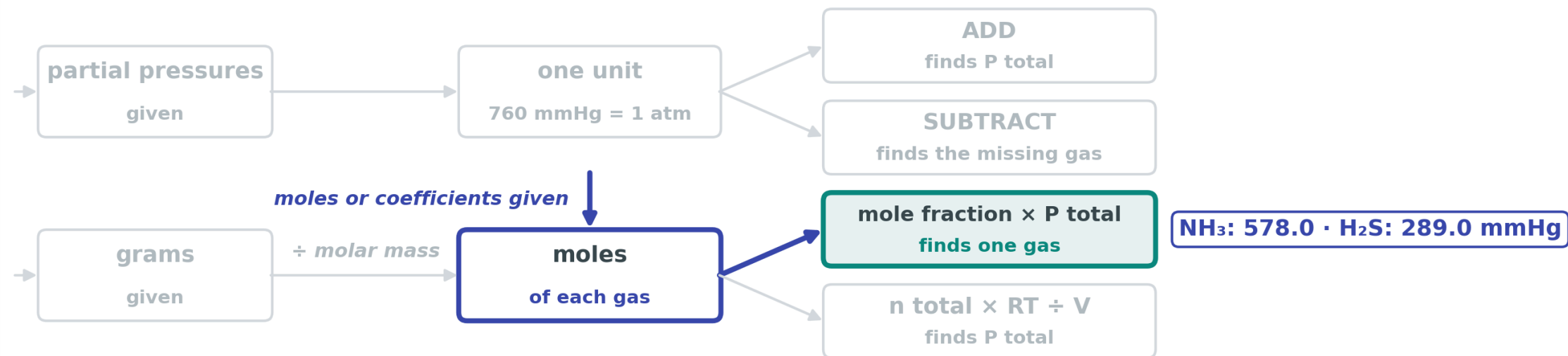
Step 4 · Check

 **Sense check:** $578.0 + 289.0 = 867.0 \text{ mmHg}$, the total. Ammonia is two-thirds of the gas molecules and carries two-thirds of the pressure. Splitting the total in half, 433.5 mmHg each, ignores the 2 : 1 ratio. ✓

✓ Worked example 3: the route on the map



coefficients: 2 mol NH_3 : 1 mol H_2S · found: NH_3 578.0 mmHg · H_2S 289.0 mmHg



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



Practice 4

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

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

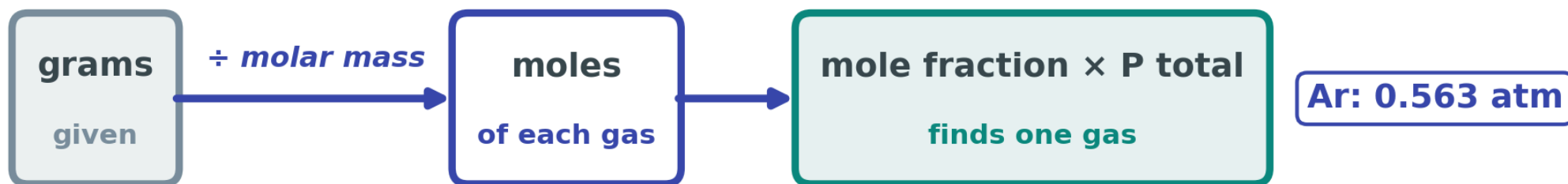
✓ Practice 4 answer: C

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

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

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

The route on the map



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



Worked example 4: total pressure from grams

$$P_{\text{total}} = n_{\text{total}}RT \div V$$

given: 0.800 g H₂ · 11.2 g N₂ · V = 5.00 L · 20.0 °C · wanted: P_{total} in atm

A 5.00 L steel flask holds **0.800 g of hydrogen** and **11.2 g of nitrogen** at **20.0 °C**. Find the total pressure in the flask.

Both gases share one volume and one temperature. List the gases, then write Dalton's law.

✓ Worked example 4: solution

$$P_{\text{total}} = n_{\text{total}}RT \div V$$

given: 0.800 g H₂ · 11.2 g N₂ · V = 5.00 L · 20.0 °C · wanted: P_{total} in atm

Step 1 · List every gas

Step 2 · Write Dalton's law

Step 3 · Find the unknown

Three moves are needed: grams to moles, add the moles, then PV = nRT.

$$0.800 \text{ g H}_2 \times \frac{1 \text{ mol}}{2.016 \text{ g}} = 0.3968 \text{ mol} \cdot 11.2 \text{ g N}_2 \times \frac{1 \text{ mol}}{28.02 \text{ g}} = 0.3997 \text{ mol}$$

$$n_{\text{total}} = 0.7965 \text{ mol} \cdot P_{\text{total}} = \frac{0.7965 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} \times 293.15 \text{ K}}{5.00 \text{ L}} = 3.83 \text{ atm}$$

 **Sense check:** One use of the ideal gas law, on the total moles of both gases, gives the total pressure. ✓

✓ Worked example 4: check

$$P_{\text{total}} = n_{\text{total}}RT \div V$$

given: 0.800 g H₂ = 0.3968 mol · 11.2 g N₂ = 0.3997 mol · 5.00 L · 293.15 K · found: P_{total} = 3.83 atm

Step 4 • Check

Each gas alone, in the same flask:

$$P_{\text{H}_2} = \frac{0.3968 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} \times 293.15 \text{ K}}{5.00 \text{ L}} = 1.91 \text{ atm}$$

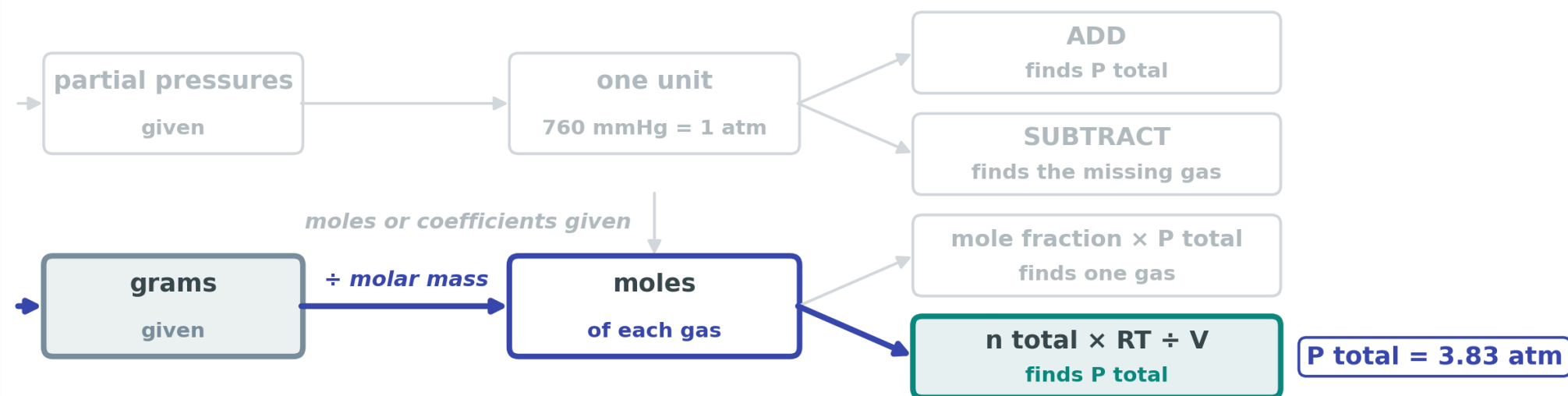
$$P_{\text{N}_2} = \frac{0.3997 \text{ mol} \times 0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} \times 293.15 \text{ K}}{5.00 \text{ L}} = 1.92 \text{ atm}$$

🧠 **Sense check:** 1.91 + 1.92 = 3.83 atm, the total. Nitrogen has 14 times the mass but about the same moles. ✓

✓ Worked example 4: the route on the map

$$P_{\text{total}} = n_{\text{total}}RT \div V$$

given: 0.800 g H₂ · 11.2 g N₂ · 5.00 L · 20.0 °C · found: P_{total} = 3.83 atm



 **Sense check:** Grams were given, so the route runs the bottom lane: molar mass to moles, then $n_{\text{total}}RT \div V$ for the whole mixture. ✓

 Practice 5

$$P_{\text{total}} = n_{\text{total}}RT \div V$$

given: 8.00 g He · 32.0 g O₂ · V = 12.0 L · 18.0 °C · wanted: P_{total}

A 12.0 L diving cylinder at **18.0 °C** holds **8.00 g of helium** and **32.0 g of oxygen**. What is the total pressure in the cylinder, in atm?

- A. 3.98
- B. 5.97
- C. 79.6
- D. 0.369

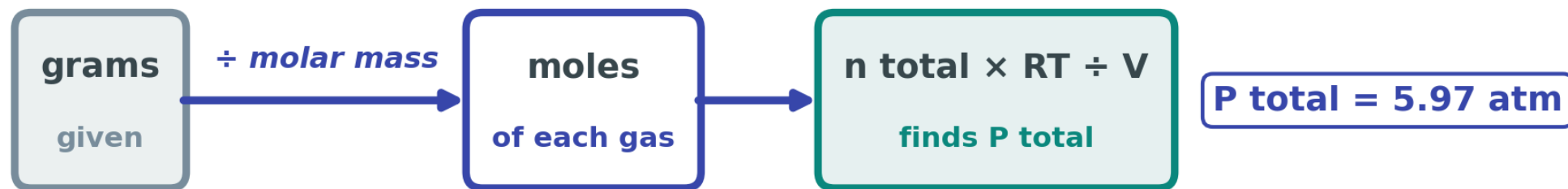
✓ Practice 5 answer: B

$$P_{\text{total}} = n_{\text{total}}RT \div V$$

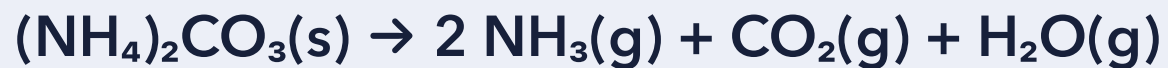
He: $8.00 \text{ g} \div 4.003 \text{ g/mol} = 1.999 \text{ mol}$ · O₂: $32.0 \text{ g} \div 32.00 \text{ g/mol} = 1.000 \text{ mol}$ · 2.999 mol total

$$P_{\text{total}} = \frac{2.999 \text{ mol} \times 0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K} \times 291.15 \text{ K}}{12.0 \text{ L}} = 5.97 \text{ atm} \rightarrow \text{answer B}$$

The route on the map



🧠 **Sense check:** Oxygen alone adds $1.000 \times 0.08206 \times 291.15 \div 12.0 = 1.99 \text{ atm}$, and $3.98 + 1.99 = 5.97 \text{ atm}$. ✓

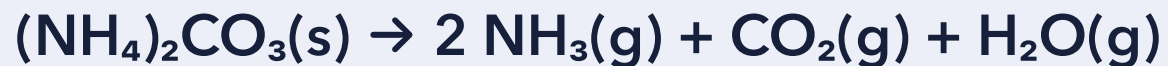
 Practice 6

given: the product gases reach $P_{\text{total}} = 884 \text{ torr}$ · wanted: $P(\text{NH}_3)$

A sample of ammonium carbonate decomposes completely in a sealed, heated flask. The product gases exert a total pressure of **884 torr**. What is the partial pressure of ammonia, in torr?

- A. 295
- B. 221
- C. 354
- D. 442

✓ Practice 6 answer: D



gases: 2 mol NH_3 + 1 mol CO_2 + 1 mol H_2O = 4 mol · $P_{\text{total}} = 884 \text{ torr}$

$$P_{\text{NH}_3} = \frac{2 \text{ mol NH}_3}{4 \text{ mol gas}} \times 884 \text{ torr} = 442 \text{ torr} \rightarrow \text{answer D}$$

 **Sense check:** Ammonia is half of the gas molecules, so it carries half of 884 torr. CO_2 and H_2O take 221 torr each: $221 + 221 + 442 = 884 \text{ torr}$. ✓

 Practice 7

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

given: 5.00 g He · 24.0 g O₂ · 84.0 g N₂ · P_{total} = 10.4 atm · wanted: P(O₂)

A trimix cylinder for a deep dive is filled with **5.00 g of helium**, **24.0 g of oxygen**, and **84.0 g of nitrogen**. The total pressure inside is **10.4 atm**. What is the partial pressure of oxygen, in atm?

- A. 1.56
- B. 3.47
- C. 2.21
- D. 2.08

✓ Practice 7 answer: A

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

$$\text{He: } 5.00 \text{ g} \div 4.003 \text{ g/mol} = 1.249 \text{ mol} \cdot \text{O}_2: 24.0 \text{ g} \div 32.00 \text{ g/mol} = 0.7500 \text{ mol} \cdot \text{N}_2: 84.0 \text{ g} \div 28.02 \text{ g/mol} = 2.998 \text{ mol} \cdot 4.997 \text{ mol total}$$

$$X_{\text{O}_2} = \frac{0.7500 \text{ mol O}_2}{4.997 \text{ mol total}} = 0.1501 \cdot P_{\text{O}_2} = 0.1501 \times 10.4 \text{ atm} = \mathbf{1.56 \text{ atm} \rightarrow \text{answer A}}$$

🧠 **Sense check:** Oxygen is 21% of the mass but only 15% of the molecules, so it carries 15% of the 10.4 atm. ✓



Practice 8

$$n_{\text{total}} = V \div 22.4 \text{ L/mol at STP}$$

given: 12.3 L at STP · 0.150 mol He · the rest N₂ · wanted: P(He)

A **12.3 L** flask at STP holds **0.150 mol of helium** mixed with nitrogen. What is the partial pressure of helium, in mmHg?

- A. 380
- B. 552
- C. 208
- D. 286

✓ Practice 8 answer: C

$$n_{\text{total}} = V \div 22.4 \text{ L/mol at STP}$$

given: 12.3 L at STP · 0.150 mol He · the rest N₂ · STP: P_{total} = 1 atm = 760 mmHg

$$12.3 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.5491 \text{ mol total}$$

$$X_{\text{He}} = \frac{0.150 \text{ mol He}}{0.5491 \text{ mol total}} = 0.2732 \cdot P_{\text{He}} = 0.2732 \times 760 \text{ mmHg}$$

= 208 mmHg → answer C

 **Sense check:** Helium is a little over a quarter of the molecules, so it carries a little over a quarter of 760 mmHg. Nitrogen takes the rest: 208 + 552 = 760 mmHg. ✓

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.

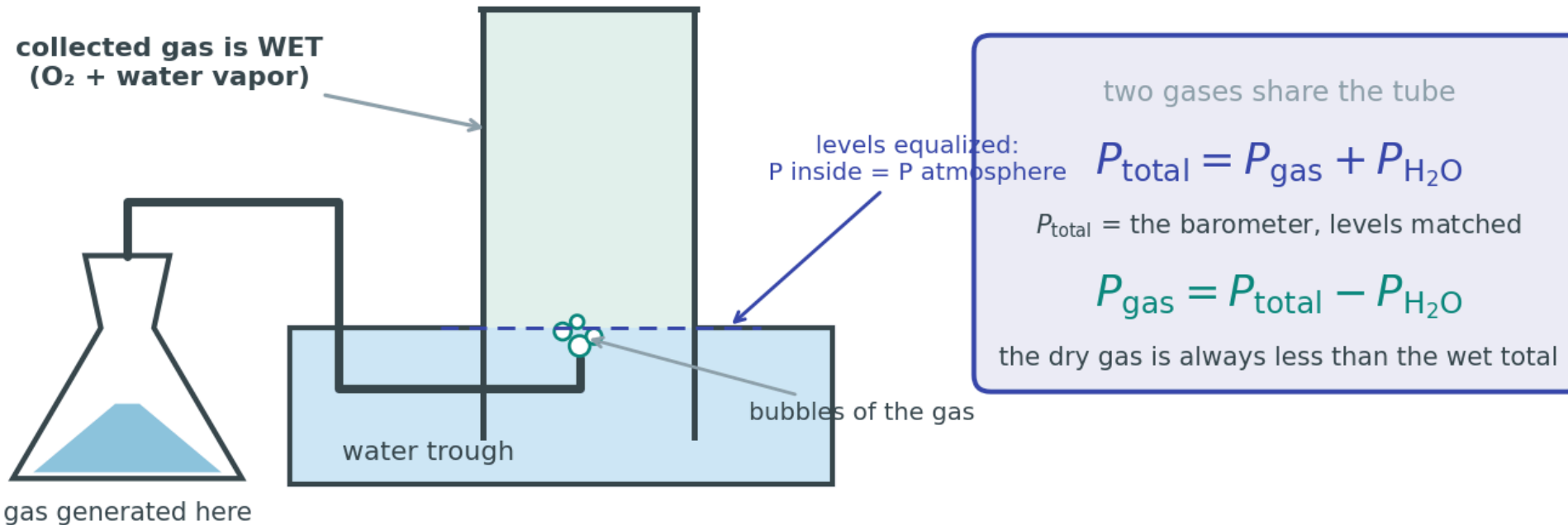
A gas bubbled up through water is a mixture too: the gas plus water vapor. Subtracting the vapor's partial pressure leaves the dry gas, and that pressure feeds the ideal gas law.

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.

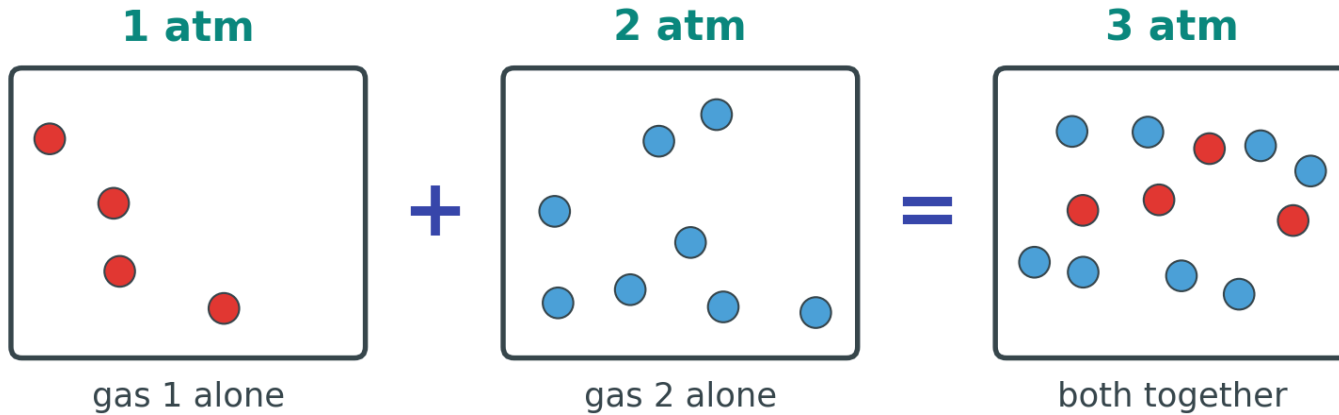
The gas comes out wet

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.

Pressures of gases in one container add



$$P_{\text{total}} = P_1 + P_2$$

same container, same temperature: 1 atm + 2 atm = 3 atm

Each gas in a mixture pushes as if it were alone, so the pressures add. This is Dalton's law. A gas collected over water always shares its container with water vapor.

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_{total} = 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.



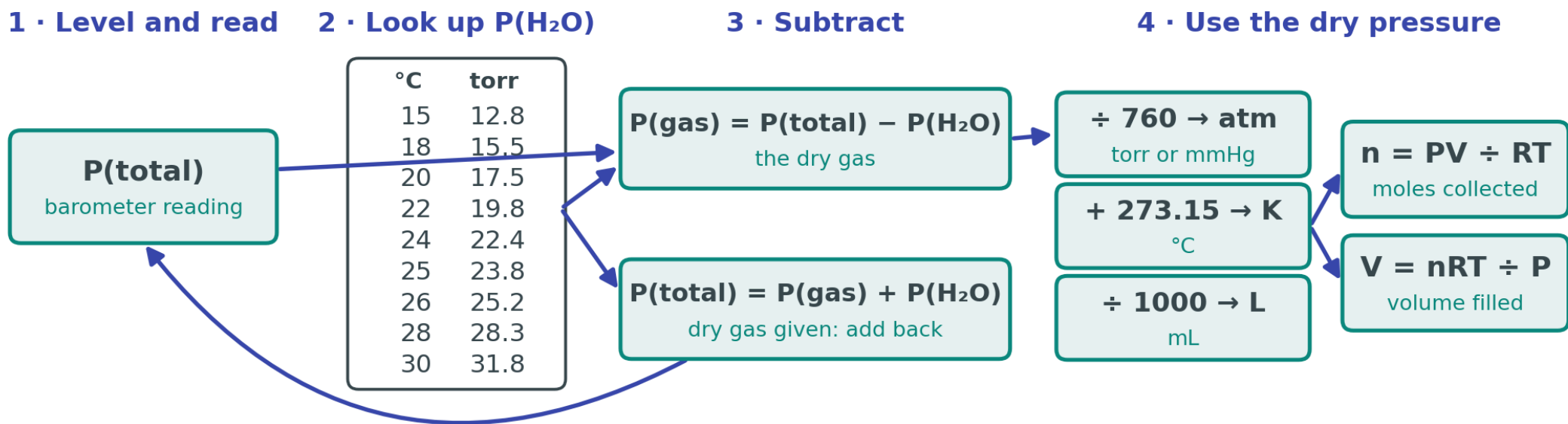
The method

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

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.

The route from the barometer to the answer



Every problem starts at the barometer and the table row. A pressure question stops after the subtraction. A moles or volume question carries the dry pressure on through step 4.



Guided example: oxygen from potassium chlorate

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

given: O₂ over water · 18 °C · barometer 745.0 torr · wanted: P(O₂) in torr

Heating potassium chlorate, KClO₃, releases oxygen. The oxygen bubbles into a bottle standing in **18 °C** water. With the levels matched, the barometer reads **745.0 torr**. Find the pressure of the dry oxygen.

Three moves: read the total, look up the vapor, subtract.

✓ Guided example: solution

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

given: O₂ over water · 18 °C · barometer 745.0 torr · wanted: P(O₂) in torr

Step 1 · Level and read

Step 2 · Look up P_{H₂O}

Step 3 · Subtract

Move 3. Take the vapor's share out of the total.

$$P_{\text{O}_2} = 745.0 \text{ torr} - 15.5 \text{ torr} = \mathbf{729.5 \text{ torr}}$$

 **Sense check:** The dry oxygen sits below the 745.0 torr wet total, as a part must. Adding gives 760.5 torr, a part larger than the whole. ✓

✓ Guided example: the route on the map

$$P_{O_2} = 745.0 \text{ torr} - 15.5 \text{ torr} = 729.5 \text{ torr}$$

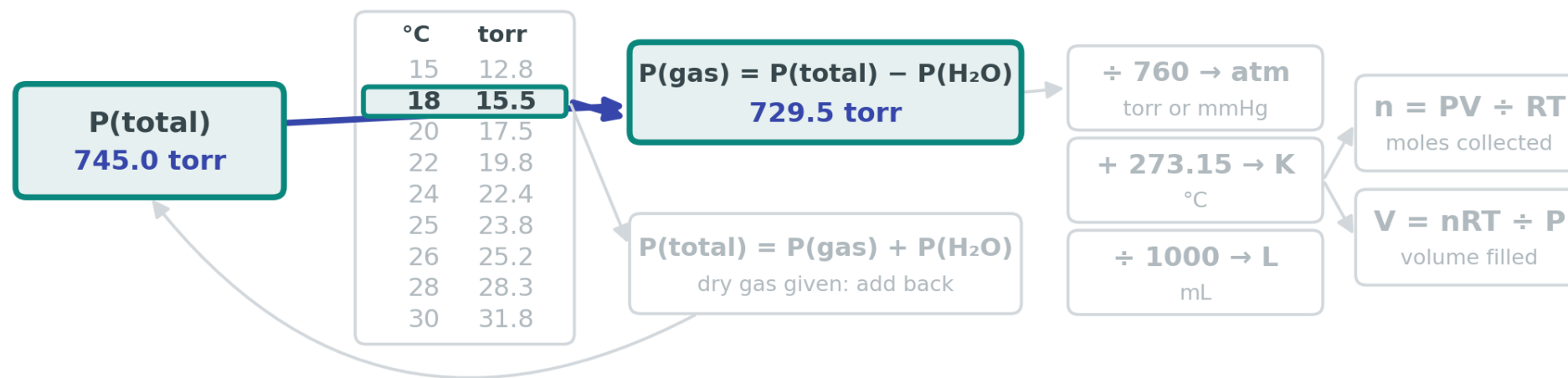
given: barometer 745.0 torr · 18 °C row: 15.5 torr · found: dry O₂ 729.5 torr

1 · Level and read

2 · Look up P(H₂O)

3 · Subtract

4 · Use the dry pressure



🧠 **Sense check:** Only a pressure was asked, so steps 1 to 3 finish the job. Step 4 stays grey. ✓



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

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

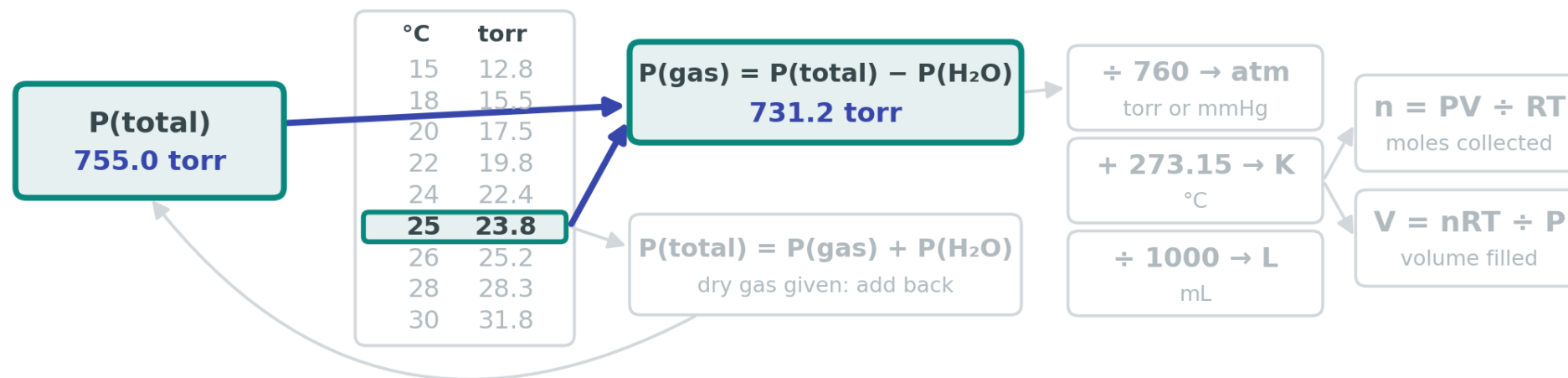
given: barometer 755.0 torr · 25 °C row: 23.8 torr · found: dry O₂ 731.2 torr

1 · Level and read

2 · Look up P(H₂O)

3 · Subtract

4 · Use the dry pressure



Sense check: The same three moves as the guided example. Only the row changed: warmer water, more vapor to subtract. ✓



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

Step 1 · Level and read

Step 2 · Look up $P_{\text{H}_2\text{O}}$

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

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

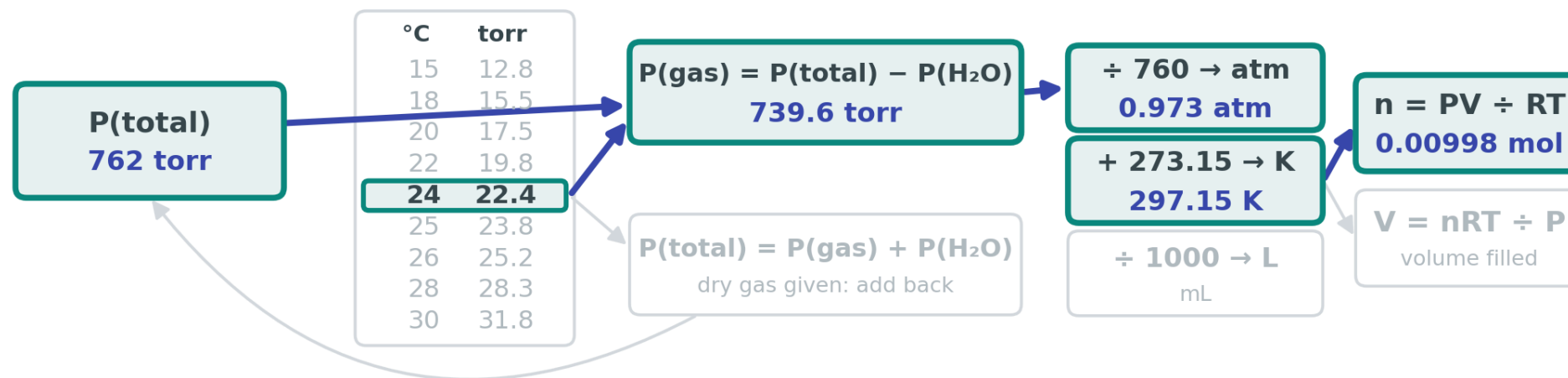
given: barometer 762 torr · 24 °C row: 22.4 torr · V = 0.250 L · found: 0.00998 mol H₂

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.

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

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

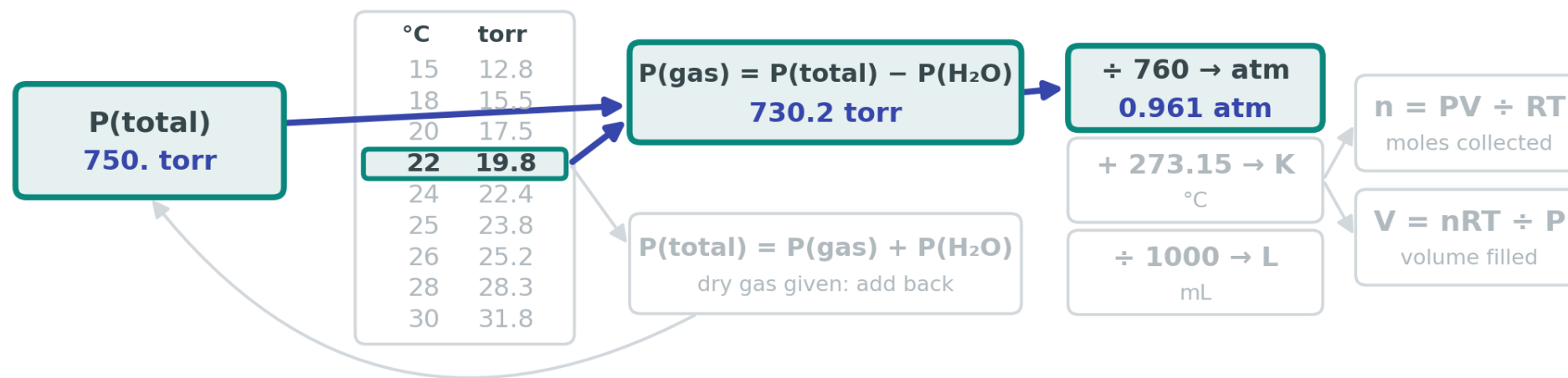
 **Sense check:** The subtraction leaves 730.2 torr of dry nitrogen, just under the 750. torr wet total. PV = nRT takes 0.961 atm, not 750 ÷ 760 = 0.987 atm. ✓

✓ Your turn: the route on the map

$$P_{N_2} = (750. \text{ torr} - 19.8 \text{ torr}) \times 1 \text{ atm} / 760 \text{ torr} = 0.961 \text{ atm}$$

given: barometer 750. torr · 22 °C row: 19.8 torr · found: dry N₂ 0.961 atm

1 · Level and read 2 · Look up P(H₂O) 3 · Subtract 4 · Use the dry pressure



🧠 **Sense check:** Step 4 went only as far as atm. With no volume or amount asked, $PV = nRT$ stays grey. ✓

⚠ 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

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

given: H_2 over water · levels matched · table: $26\text{ }^\circ\text{C} \rightarrow 25.2\text{ torr}$ · wanted: $P(\text{H}_2)$ in torr

Magnesium ribbon in acid gives off hydrogen, caught in an inverted tube over $26\text{ }^\circ\text{C}$ water. The barometer reads **774.0 torr**. What pressure, in torr, does the dry hydrogen exert?

- A. 774.0
- B. 748.8
- C. 734.8
- D. 799.2

✓ Practice 1 answer: B

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

levels matched: $P_{\text{total}} = \text{barometer } 774.0 \text{ torr} \cdot 26^\circ\text{C row: } 25.2 \text{ torr}$

$$P_{\text{H}_2} = 774.0 \text{ torr} - 25.2 \text{ torr} = \mathbf{748.8 \text{ torr} \rightarrow \text{answer B}}$$

A skipped the vapor: 774.0 torr is hydrogen and water vapor together. C subtracted from the standard atmosphere, not today's reading: $760 - 25.2 = 734.8 \text{ torr}$. D added the vapor: $774.0 + 25.2 = 799.2 \text{ torr}$, a part larger than the whole.

🧠 **Sense check:** The dry hydrogen sits exactly one vapor share, 25.2 torr, under the wet total. ✓

✓ Practice 1: the route on the map

$$P_{\text{H}_2} = 774.0 \text{ torr} - 25.2 \text{ torr} = 748.8 \text{ torr}$$

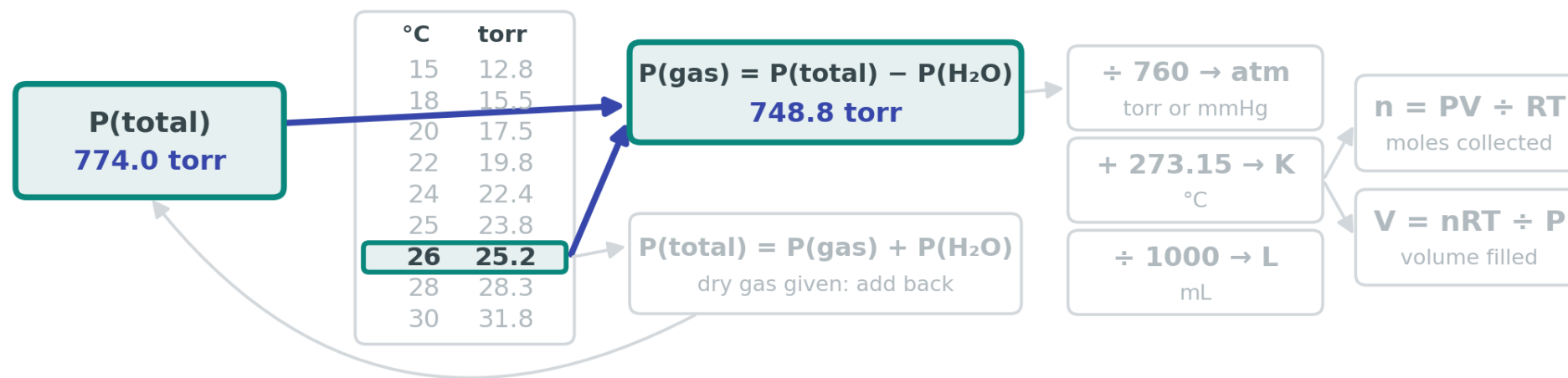
given: barometer 774.0 torr · 26 °C row: 25.2 torr · found: dry H₂ 748.8 torr

1 · Level and read

2 · Look up P(H₂O)

3 · Subtract

4 · Use the dry pressure



Sense check: A pressure question: steps 1 to 3, the same path as the guided example. ✓



Practice 2

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

given: CH_4 over water · 25 °C · levels matched · $P_{\text{CH}_4} = 727.4$ 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 2 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: the route on the map

$$P_{\text{atm}} = 727.4 \text{ torr} + 23.8 \text{ torr} = 751.2 \text{ torr}$$

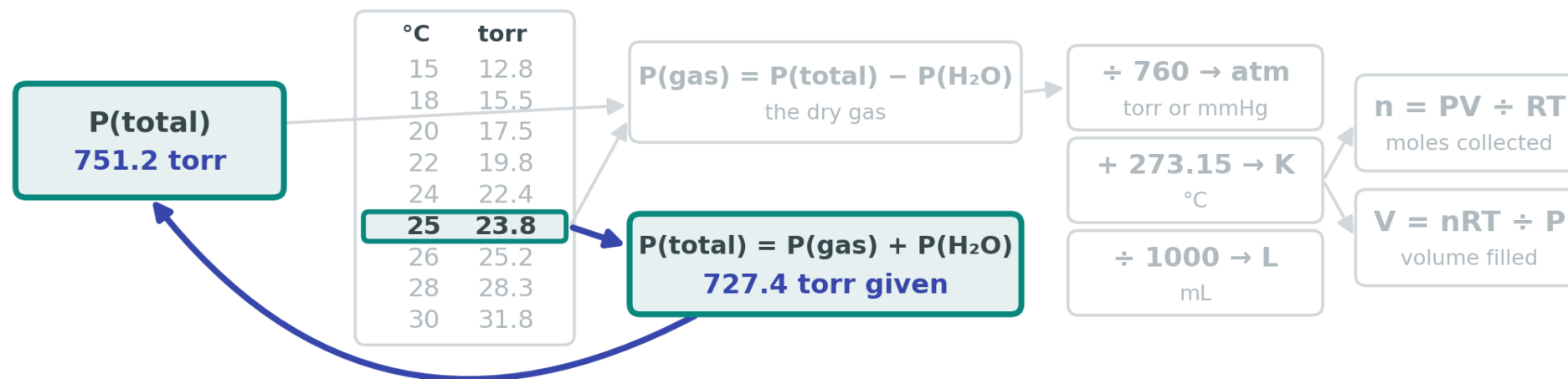
given: dry CH_4 727.4 torr · 25 °C row: 23.8 torr · found: the barometer, 751.2 torr

1 · Level and read

2 · Look up $P(\text{H}_2\text{O})$

3 · Subtract

4 · Use the dry pressure



Sense check: The dry pressure was given, so step 3 ran backward: add the vapor back, and the sum is the barometer reading. ✓

 Practice 3

$$P_{\text{gas}} = (P_{\text{total}} - P_{\text{H}_2\text{O}}) \times 1 \text{ atm} / 760 \text{ torr}$$

given: N_2 over water · levels matched · table: $28^\circ\text{C} \rightarrow 28.3 \text{ torr}$ · wanted: $P(\text{N}_2)$ in atm

Warming an ammonium nitrite solution releases nitrogen, collected over 28°C water. The barometer reads **739 torr**. What is the pressure of the dry nitrogen, in atm?

- A. 1.01
- B. 0.0372
- C. 0.935
- D. 0.972

✓ Practice 3 answer: C

$$P_{\text{gas}} = (P_{\text{total}} - P_{\text{H}_2\text{O}}) \times 1 \text{ atm} / 760 \text{ torr}$$

barometer 739 torr · 28 °C row: 28.3 torr · wanted: P(N₂) in atm

$$P_{\text{N}_2} = 739 \text{ torr} - 28.3 \text{ torr} = 710.7 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.935 \text{ atm} \rightarrow \text{answer C}$$

A added the vapor: $(739 + 28.3) \div 760 = 1.01 \text{ atm}$. B converted the vapor instead of the gas: $28.3 \div 760 = 0.0372 \text{ atm}$ is water's share. D skipped the subtraction: $739 \div 760 = 0.972 \text{ atm}$ counts the vapor as nitrogen.

 **Sense check:** The wet 0.972 atm minus the vapor's 0.0372 atm leaves 0.935 atm. ✓

✓ Practice 3: the route on the map

$$P_{N_2} = (739 \text{ torr} - 28.3 \text{ torr}) \times 1 \text{ atm} / 760 \text{ torr} = 0.935 \text{ atm}$$

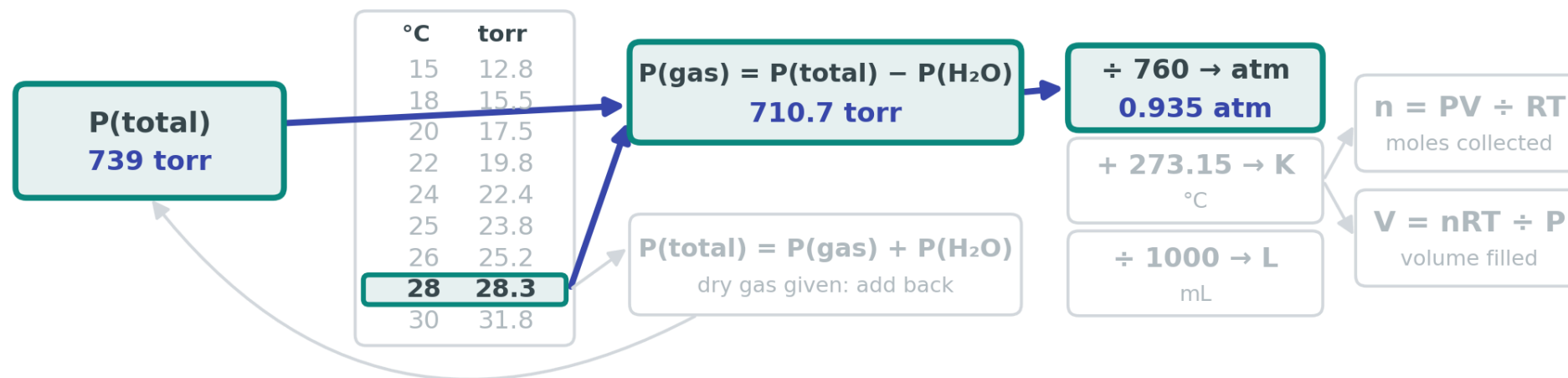
given: barometer 739 torr · 28 °C row: 28.3 torr · found: dry N₂ 0.935 atm

1 · Level and read

2 · Look up P(H₂O)

3 · Subtract

4 · Use the dry pressure



Sense check: One unit match past the subtraction, torr to atm, and no gas law. The same path as the faded nitrogen problem. ✓

 Practice 4

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

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 4 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. ✓

✓ Practice 4: the route on the map

$$V = nRT / P_{\text{gas}}$$

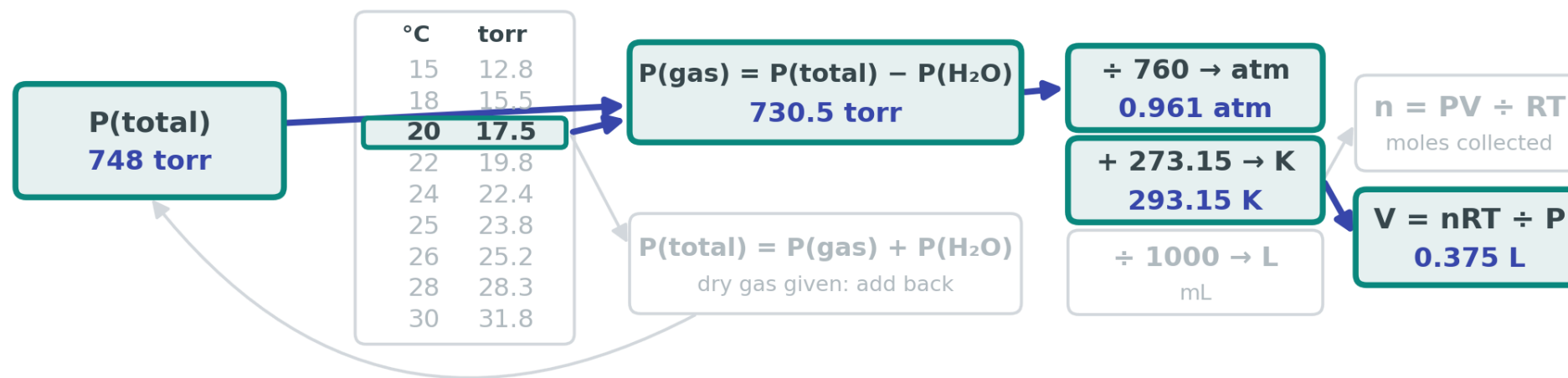
given: 0.0150 mol O₂ · barometer 748 torr · 20 °C row: 17.5 torr · found: 0.375 L

1 · Level and read

2 · Look up P(H₂O)

3 · Subtract

4 · Use the dry pressure



🧠 **Sense check:** The same four steps as a moles problem. The gas law was solved for V instead of n. ✓



Practice 5

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

given: H_2 over water · levels matched · table: $15\text{ }^\circ\text{C} \rightarrow 12.8\text{ torr}$ · $20\text{ }^\circ\text{C} \rightarrow 17.5\text{ torr}$ · $25\text{ }^\circ\text{C} \rightarrow 23.8\text{ torr}$ ·
wanted: mol H_2

Zinc dissolving in hydrochloric acid gives off hydrogen. Over **$15\text{ }^\circ\text{C}$** water, **385 mL** of it fills the tube while the room's pressure is **742 mmHg** . What amount of H_2 , in moles, was captured?

- A. 0.0156
- B. 11.9
- C. 0.300
- D. 0.0159
- E. 15.6

✓ Practice 5 answer: A

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

$$742 \text{ mmHg} \cdot 15 \text{ }^{\circ}\text{C row: } 12.8 \text{ torr} = 12.8 \text{ mmHg} \cdot 385 \text{ mL} \cdot T = 15 + 273.15 = 288.15 \text{ K}$$

$$P_{\text{H}_2} = 742 \text{ mmHg} - 12.8 \text{ mmHg} = 729.2 \text{ mmHg} = \mathbf{0.959 \text{ atm}}$$

$$V = 385 \text{ mL} \div 1000 = \mathbf{0.385 \text{ L}}$$

$$n = \frac{0.959 \text{ atm} \times 0.385 \text{ L}}{0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} \times 288.15 \text{ K}} = \mathbf{0.0156 \text{ mol} \rightarrow \text{answer A}}$$

B kept mmHg: $729.2 \times 0.385 \div 23.6 = 11.9$. C used $15 \text{ }^{\circ}\text{C}$ for T: 0.300. D used the wet 742 mmHg, 0.976 atm: 0.0159. E kept 385 mL as 385: 15.6.

 **Sense check:** Near room temperature and 1 atm a mole fills about 24 L, so 0.385 L holds about $0.385 \div 24 = 0.016$ mol. ✓

✓ Practice 5: the route on the map

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

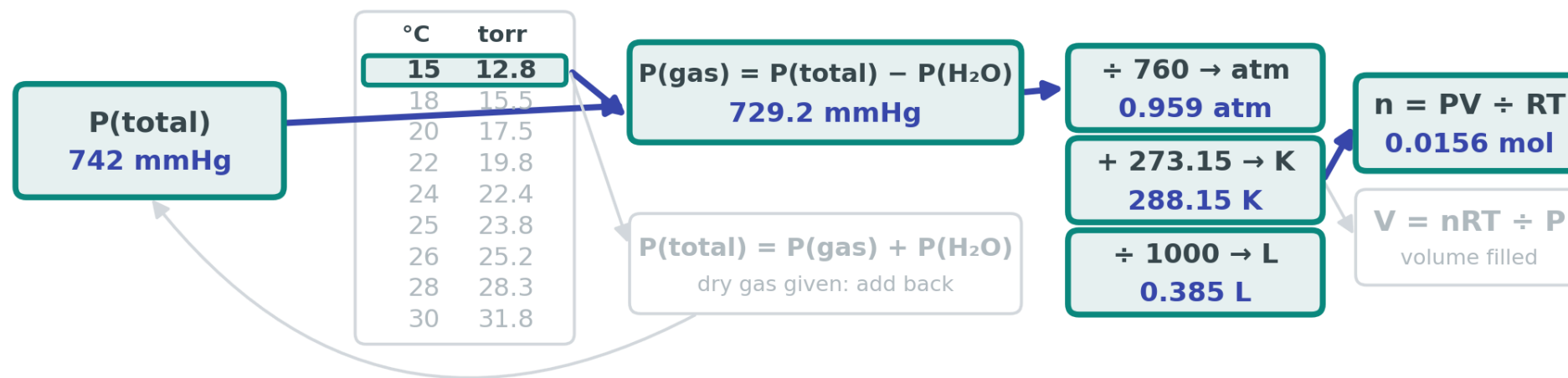
given: 742 mmHg · 15 °C row: 12.8 torr · 385 mL · found: 0.0156 mol H₂

1 · Level and read

2 · Look up P(H₂O)

3 · Subtract

4 · Use the dry pressure



🧠 **Sense check:** Every unit match in step 4 was needed: mmHg to atm, °C to K, and mL to L. ✓

 Practice 6

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

given: propane over water · levels matched · table: $27\text{ }^{\circ}\text{C} \rightarrow 26.7\text{ torr} \cdot 1\text{ atm} = 760\text{ torr} = 101.325\text{ kPa} \cdot$

wanted: $P(\text{propane})$ in torr

Propane from a camp-stove canister bubbles into an inverted bottle over **$27\text{ }^{\circ}\text{C}$** water. With the levels matched, the barometer reads **100.4 kPa** . What pressure, in torr, does the dry propane exert?

- A. 753.1
- B. 779.8
- C. 726.4
- D. 73.7

✓ Practice 6 answer: C

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

barometer 100.4 kPa · 27 °C row: 26.7 torr · two units: match them before subtracting

$$P_{\text{total}} = 100.4 \text{ kPa} \times \frac{760 \text{ torr}}{101.325 \text{ kPa}} = 753.1 \text{ torr}$$

$$P_{\text{propane}} = 753.1 \text{ torr} - 26.7 \text{ torr} = 726.4 \text{ torr} \rightarrow \text{answer C}$$

A converted the barometer but skipped the vapor: 753.1 torr is propane and water vapor together. B added the vapor: $753.1 + 26.7 = 779.8$ torr, a part larger than the whole. D subtracted across units: $100.4 - 26.7 = 73.7$ takes torr away from kPa.

 **Sense check:** 100.4 kPa is just under 1 atm, so the wet total is just under 760 torr. The dry gas sits one vapor share below it. ✓

 Practice 7

$$P_1V_1 / T_1 = P_2V_2 / T_2, \text{ with } P_1 = P_{\text{total}} - P_{\text{H}_2\text{O}}$$

given: O₂ over water · levels matched · table: 21 °C → 18.7 torr · wanted: V of the dry O₂ at STP, in mL

Pondweed under a lamp gives off oxygen, caught in a leveled tube over **21 °C** water. The tube holds **24.5 mL** while the barometer reads **751 torr**. What volume, in mL, would the dry oxygen occupy at STP?

- A. 23.6
- B. 21.9
- C. 25.4
- D. 22.5

✓ Practice 7 answer: B

$$V_2 = V_1 \times (P_1 / P_2) \times (T_2 / T_1)$$

collected: 24.5 mL · 751 torr wet · 21 °C row: 18.7 torr · STP: 760 torr and 273.15 K

$$P_1 = 751 \text{ torr} - 18.7 \text{ torr} = \mathbf{732.3 \text{ torr}} \cdot T_1 = 21 + 273.15 = \mathbf{294.15 \text{ K}}$$

$$V_2 = 24.5 \text{ mL} \times \frac{732.3 \text{ torr}}{760 \text{ torr}} \times \frac{273.15 \text{ K}}{294.15 \text{ K}} = \mathbf{21.9 \text{ mL} \rightarrow \text{answer B}}$$

A flipped the pressure ratio: $24.5 \times 760/732.3 \times 273.15/294.15 = 23.6 \text{ mL}$. C flipped the temperature ratio: $24.5 \times 732.3/760 \times 294.15/273.15 = 25.4 \text{ mL}$. D used the wet 751 torr: $24.5 \times 751/760 \times 273.15/294.15 = 22.5 \text{ mL}$, counting the vapor as oxygen.

 **Sense check:** At STP the gas is squeezed harder (760 vs 732.3 torr) and colder (273 vs 294 K). Both shrink it, so V_2 falls below 24.5 mL. ✓

 Practice 8

$$n = P_{\text{gas}}V / (RT), \text{ then mass} = n \times M$$

given: C_4H_{10} over water · levels matched · table: $19^\circ\text{C} \rightarrow 16.5 \text{ torr}$ · wanted: g C_4H_{10}

Butane, C_4H_{10} , from a lighter is bubbled into a leveled **0.425 L** bottle over **19°C** water. The barometer reads **744 torr**. What mass of butane, in grams, was collected?

- A. 1.01
- B. 15.2
- C. 750
- D. 0.986
- E. 1.03

✓ Practice 8 solution: moles of butane

$$n = P_{\text{gas}}V / (RT), \text{ then mass} = n \times M$$

$$744 \text{ torr} \cdot 19^\circ\text{C row: } 16.5 \text{ torr} \cdot 0.425 \text{ L} \cdot \text{wanted: g C}_4\text{H}_{10}$$

$$P_{\text{C}_4\text{H}_{10}} = 744 \text{ torr} - 16.5 \text{ torr} = 727.5 \text{ torr} = \mathbf{0.957 \text{ atm}} \cdot T = 19 + 273.15 = \mathbf{292.15 \text{ K}}$$

$$n = \frac{0.957 \text{ atm} \times 0.425 \text{ L}}{0.08206 \text{ L atm/mol} \cdot \text{K} \times 292.15 \text{ K}} = \mathbf{0.01697 \text{ mol}}$$

🧠 **Sense check:** The dry 727.5 torr, not the wet 744 torr, goes into $PV = nRT$. ✓

✓ Practice 8 answer: D

$$\text{mass} = n \times M$$

$$n = 0.01697 \text{ mol C}_4\text{H}_{10} \cdot \text{C}_4\text{H}_{10}: 4(12.01) + 10(1.008) = 58.12 \text{ g/mol}$$

$$0.01697 \text{ mol} \times \frac{58.12 \text{ g}}{1 \text{ mol}} = 0.986 \text{ g} \rightarrow \text{answer D}$$

A used the wet 744 torr: 1.01 g. E added the vapor, 760.5 torr: 1.03 g. B used 19 °C for T: 15.2 g. C kept torr: $727.5 \times 0.425 \div 23.97 \times 58.12 = 750$.

 **Sense check:** Near room temperature and 1 atm a mole fills about 24 L, so 0.425 L holds about 0.018 mol, near 1 g of butane. The dry gas sits below 1 atm, so a little less. ✓

Check yourself

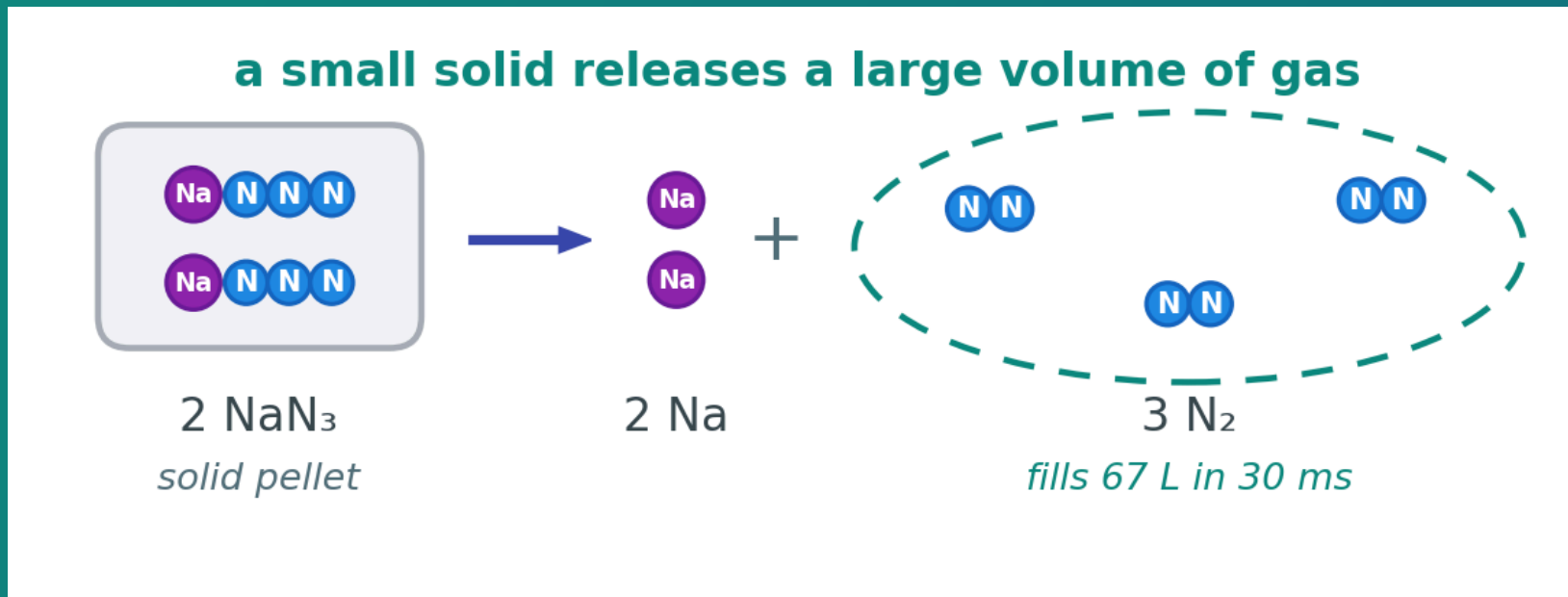
1. Hydrogen is collected over water at 30 °C in a leveled tube. The barometer reads 770 torr and water vapor at 30 °C is 31.8 torr. Find the dry H₂ pressure, and state which pressure you would put into $PV = nRT$.
2. Why does using the barometer reading directly always overcount the moles of gas collected?

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

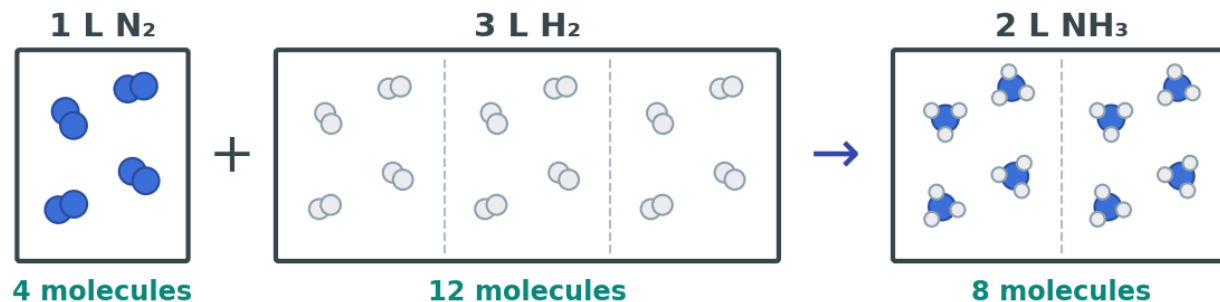
Bring a gas volume into the mole bridge (as moles at STP through 22.4 L/mol, or off STP through $PV = nRT$), then cross the mole ratio to reach grams or a second gas volume; for two gases at the same temperature and pressure, read the coefficients directly in liters.

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.

Coefficients as gas volumes



1 L : 3 L : 2 L at the same T and P · 4 L of gas react, 2 L form

At the same temperature and pressure, equal gas volumes hold equal numbers of molecules. So gas coefficients also read in liters. How many liters of H₂ react with 5.0 L of N₂?

$$5.0 \text{ L N}_2 \times \frac{3 \text{ L H}_2}{1 \text{ L N}_2} = 15 \text{ L H}_2$$

A gas volume counts moles

1 mol of any gas = 22.4 L at STP

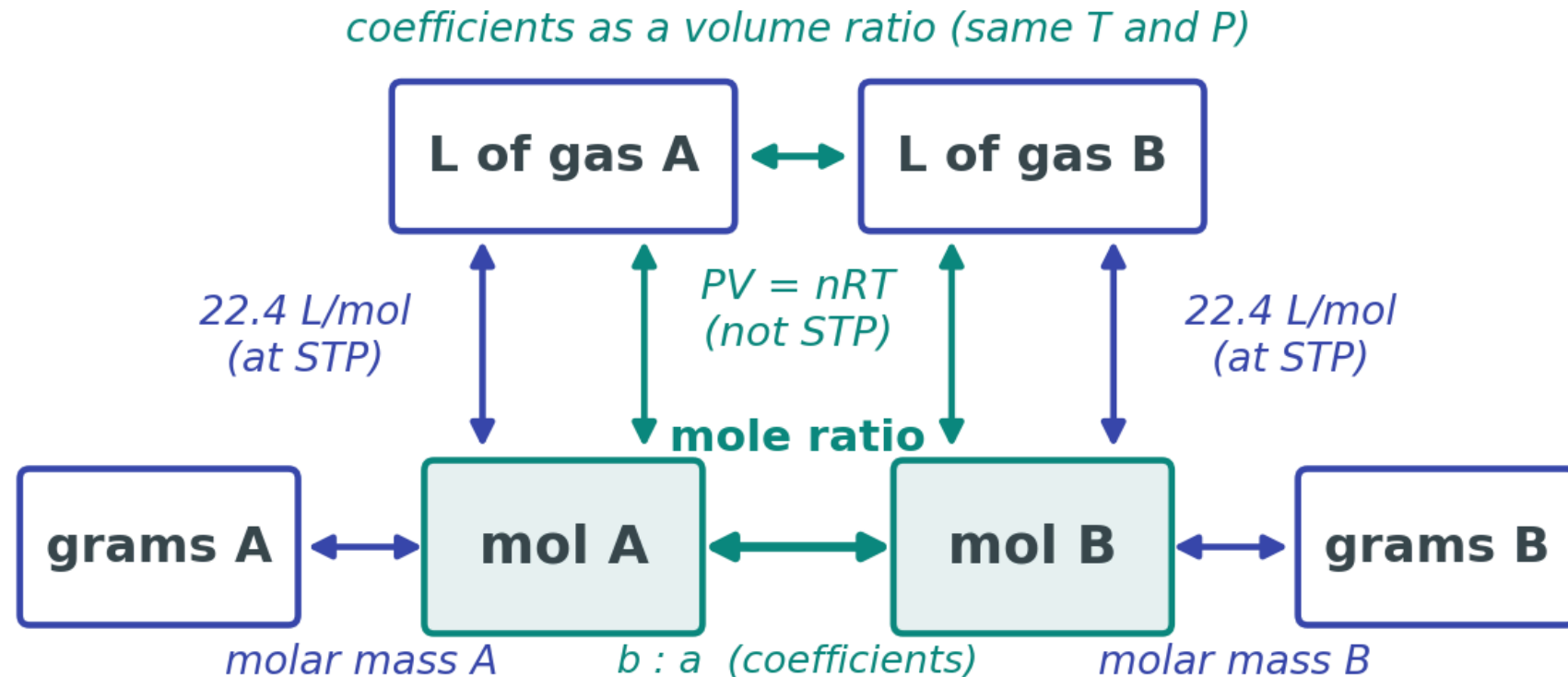
STP: 0 °C (273.15 K), 1 atm · every gas, same volume per mole

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 enters through 22.4 L/mol at STP, or $PV = nRT$ otherwise. Grams still cross the mole bridge. Two gases at the same T and P need only the coefficients, read in liters.

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.



Guided example: oxygen from mercury(II) oxide

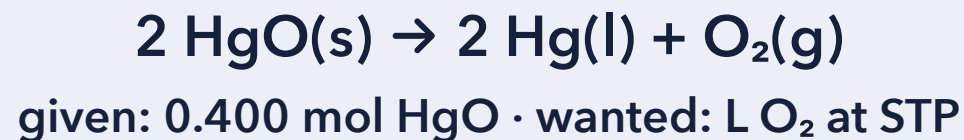


given: 0.400 mol HgO · wanted: L O₂ at STP

Heating mercury(II) oxide releases oxygen gas, one of the first ways oxygen was isolated.
0.400 mol of HgO decomposes. What volume of O₂ forms at STP?

Check each method step in turn. One is already done.

✓ Guided example: solution



Step 1 · To moles

Step 2 · Moles → moles

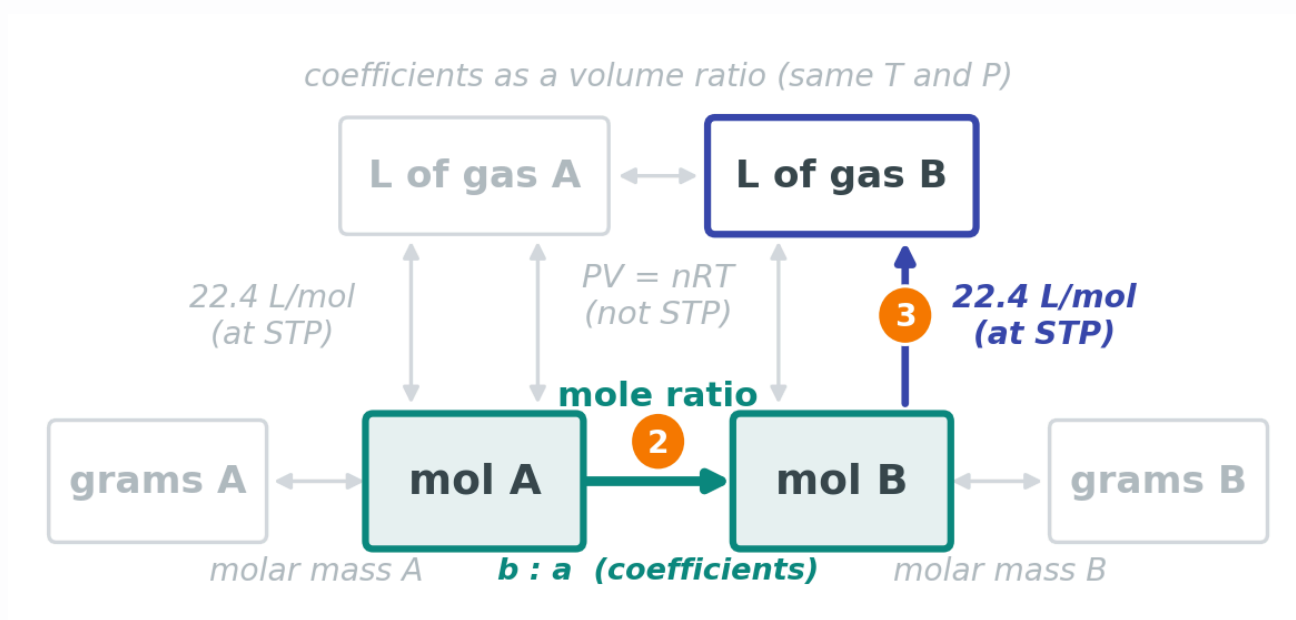
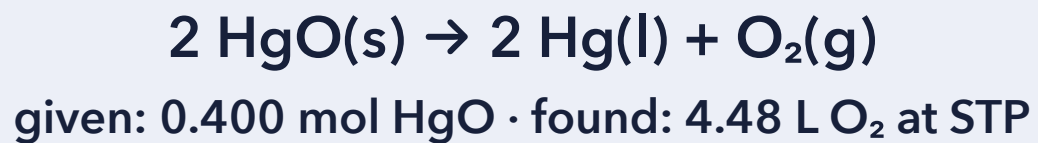
$$0.400 \cancel{\text{ mol HgO}} \times \frac{1 \text{ mol O}_2}{2 \cancel{\text{ mol HgO}}} = 0.200 \text{ mol O}_2$$

Step 3 · To the wanted unit

$$0.400 \cancel{\text{ mol HgO}} \times \frac{1 \cancel{\text{ mol O}_2}}{2 \cancel{\text{ mol HgO}}} \times \frac{22.4 \text{ L O}_2}{1 \cancel{\text{ mol O}_2}} = 4.48 \text{ L O}_2$$

 **Sense check:** Two HgO release one O₂, so 0.400 mol HgO gives 0.200 mol O₂. A fifth of a mole at 22.4 L per mole is 4.48 L. ✓

✓ Guided example: the route on the map



🧠 **Sense check:** A is HgO, B is O₂. The given was already moles, so the route starts at mol A: arrow 2 is the mole ratio, arrow 3 the molar volume. ✓

 Practice 1

Nitric oxide from car exhaust reacts with oxygen in the air. When **4.50 L of NO** reacts completely, what volume of O_2 , in liters, is consumed at the same temperature and pressure?

- A. 9.00
- B. 4.50
- C. 2.25
- D. 0.100

✓ Practice 1 answer: C



given: 4.50 L NO · wanted: L O₂ · same T and P: 2 L NO react with 1 L O₂

$$4.50 \text{ L NO} \times \frac{1 \text{ L O}_2}{2 \text{ L NO}} = 2.25 \text{ L O}_2 \rightarrow \text{answer C}$$

A inverted the ratio: $4.50 \times (2/1) = 9.00$. B assumed a 1 : 1 volume ratio: 4.50 L. D divided by 22.4: $2.25 \div 22.4 = 0.100$ would be moles, but the stem never says STP and the question asks for liters.

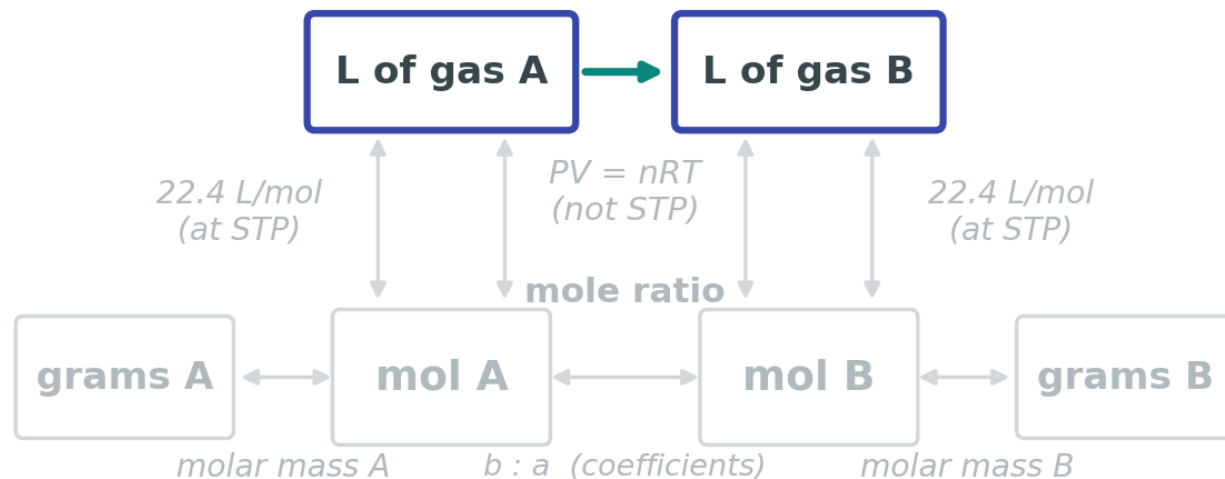
 **Sense check:** Two NO molecules use one O₂, so the O₂ volume is half the NO volume: 2.25 L. ✓

✓ Practice 1: the route on the map



given: 4.50 L NO · found: 2.25 L O₂ at the same T and P

coefficients as a volume ratio (same T and P)



 **Sense check:** Both substances are gases at the same T and P, so one arrow does it: the coefficients as a volume ratio. No trip through moles. ✓



Practice 2



A camp stove burns **5.60 L of propane**, measured at STP. How many moles of H_2O form?

- A. 0.750
- B. 1.00
- C. 0.0625
- D. 0.250

✓ Practice 2 answer: B



given: 5.60 L C₃H₈ at STP · wanted: mol H₂O

$$5.60 \text{ L C}_3\text{H}_8 \times \frac{1 \text{ mol C}_3\text{H}_8}{22.4 \text{ L C}_3\text{H}_8} \times \frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol C}_3\text{H}_8} = 1.00 \text{ mol H}_2\text{O} \rightarrow \text{answer B}$$

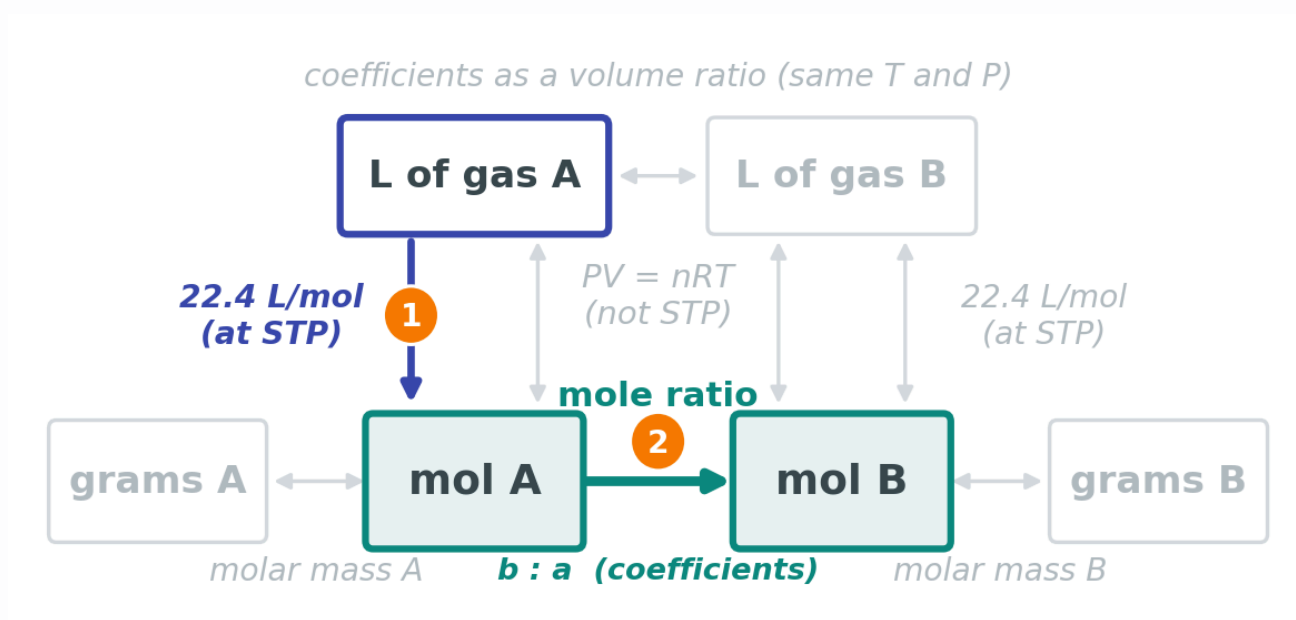
A read CO₂'s coefficient: $0.250 \times (3/1) = 0.750$. C inverted the ratio: $0.250 \times (1/4) = 0.0625$. D skipped the mole ratio: $5.60 \div 22.4 = 0.250$ mol is the propane, not the water.

 **Sense check:** 5.60 L at STP is a quarter mole of propane. Each propane makes four H₂O, so a quarter mole makes 1.00 mol. ✓

✓ Practice 2: the route on the map



given: 5.60 L C₃H₈ at STP · found: 1.00 mol H₂O



🧠 **Sense check:** A is propane, B is water. Moles were wanted, so the route stops at mol B after two arrows. D stopped one arrow early, at mol A. ✓



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

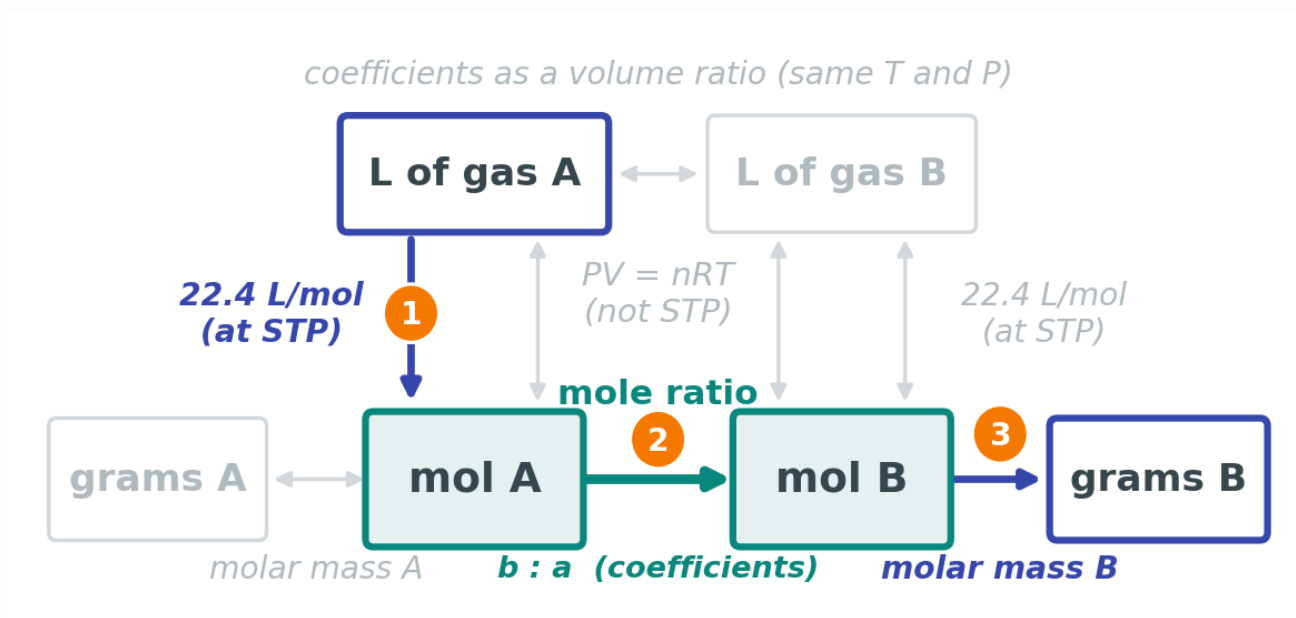
$$33.6 \cancel{\text{L H}_2} \times \frac{1 \cancel{\text{mol H}_2}}{22.4 \cancel{\text{L H}_2}} \times \frac{2 \cancel{\text{mol NH}_3}}{3 \cancel{\text{mol H}_2}} \times \frac{17.03 \text{ g NH}_3}{1 \cancel{\text{mol NH}_3}} = 17.0 \text{ g NH}_3$$

 **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

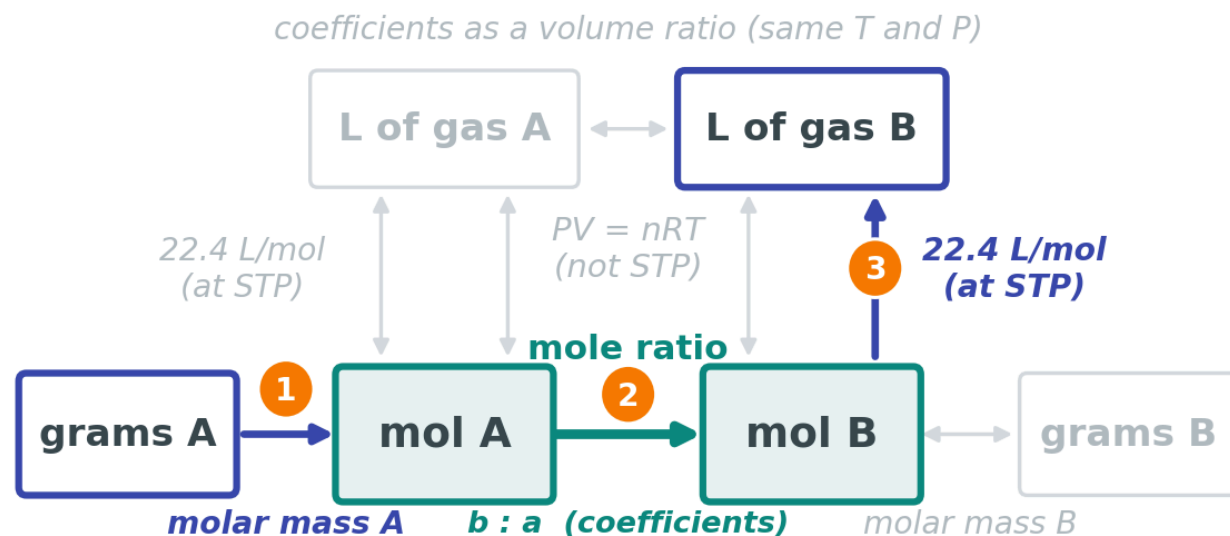
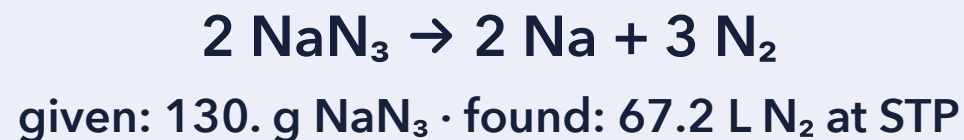
Step 2 · Moles → moles

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 3

molar mass Al 26.98 g/mol

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 3 answer: D



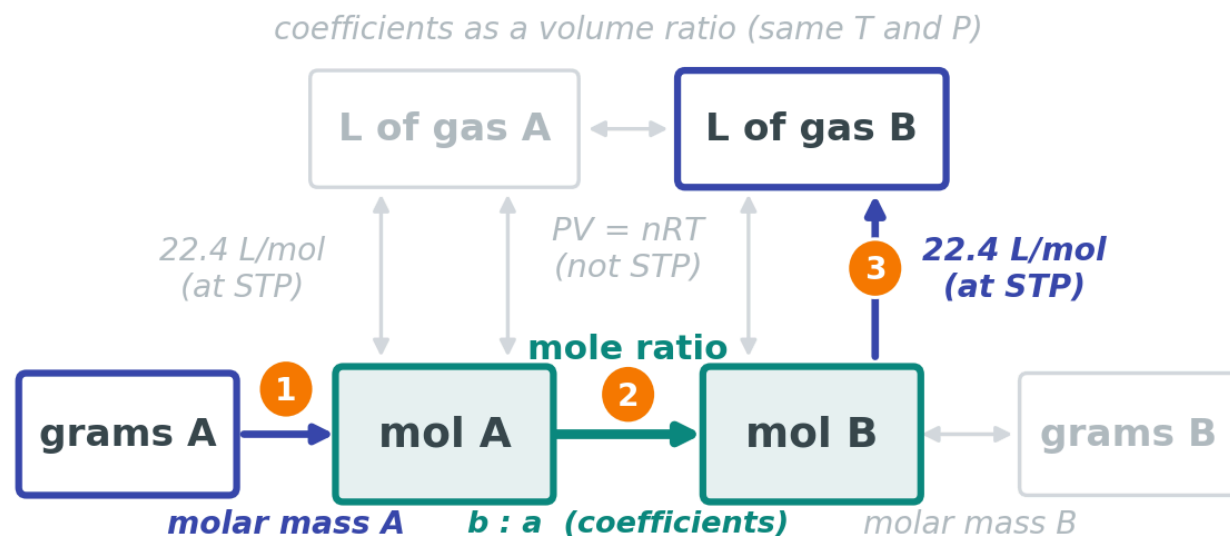
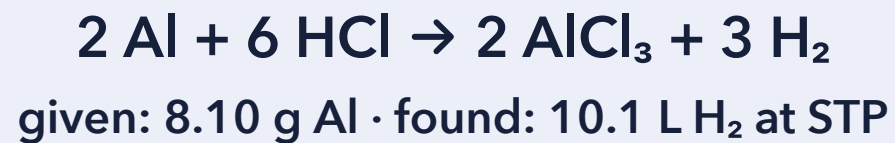
given: 8.10 g Al · wanted: L H₂ at STP

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

A skipped the mole ratio: $8.10/26.98 \times 22.4 = 6.72$. B stopped at moles: $8.10/26.98 \times 3/2 = 0.450$ 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 ✓

✓ Practice 3: the route on the map



🧠 **Sense check:** A is Al, B is H₂. Choice A skipped arrow 2, the mole ratio. B stopped before arrow 3. C ran arrow 2 upside down. ✓

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):

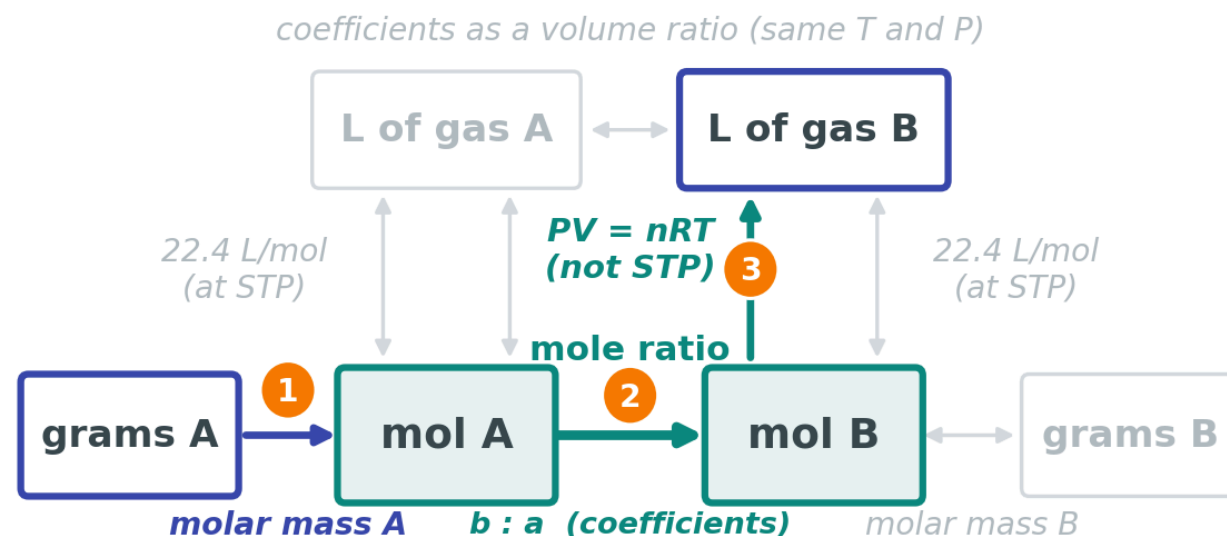
$$V = \frac{nRT}{P} = \frac{(0.250)(0.08206)(300.15)}{1.00} = \mathbf{6.16 \text{ L CO}_2}$$

🧠 **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



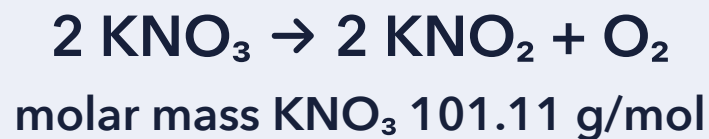
🧠 **Sense check:** Arrows 1 and 2 match an STP problem. Off STP, arrow 3 is $PV = nRT$, not 22.4 L/mol. ✓

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 4

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 4 answer: A



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

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

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

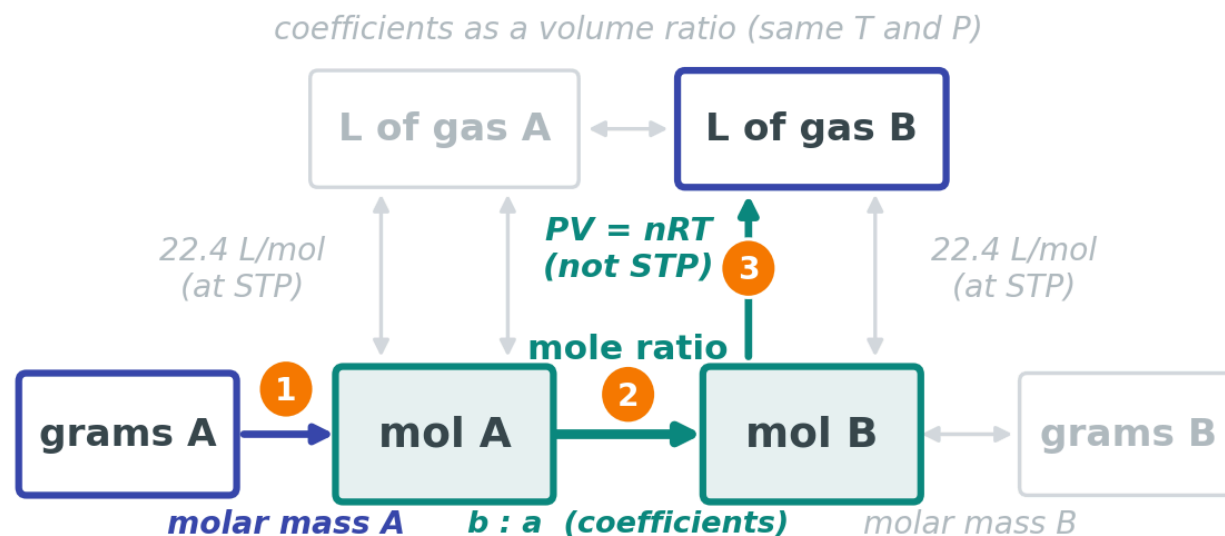
B used 22.4 L/mol off STP: $0.198 \times 22.4 = 4.43$. C stopped at moles: 0.198 mol O_2 is the amount, not the volume it fills. D forgot to convert to kelvin: $(0.198)(0.08206)(127) = 2.06$.

🧠 **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. ✓

✓ Practice 4: the route on the map



given: 40.0 g KNO_3 · found: 6.50 L O_2 at 127 °C, 1.00 atm



🧠 **Sense check:** Off STP, arrow 3 is $PV = nRT$. Choice B took the 22.4 L/mol arrow instead; C stopped before arrow 3.

✓



Practice 5



molar mass KClO_3 122.55 g/mol

A student heats potassium chlorate and collects **672 mL of O_2** , measured at STP. What mass of KClO_3 , in grams, decomposed?

- A. 3.68
- B. 0.0200
- C. 5.51
- D. 2.45
- E. 2.45×10^3

✓ Practice 5 answer: D



given: 672 mL O₂ at STP = 0.672 L · wanted: g KClO₃ (122.55 g/mol)

$$0.672 \text{ L O}_2 \times \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} \times \frac{2 \text{ mol KClO}_3}{3 \text{ mol O}_2} \times \frac{122.55 \text{ g KClO}_3}{1 \text{ mol KClO}_3} = 2.45 \text{ g} \rightarrow \text{answer D}$$

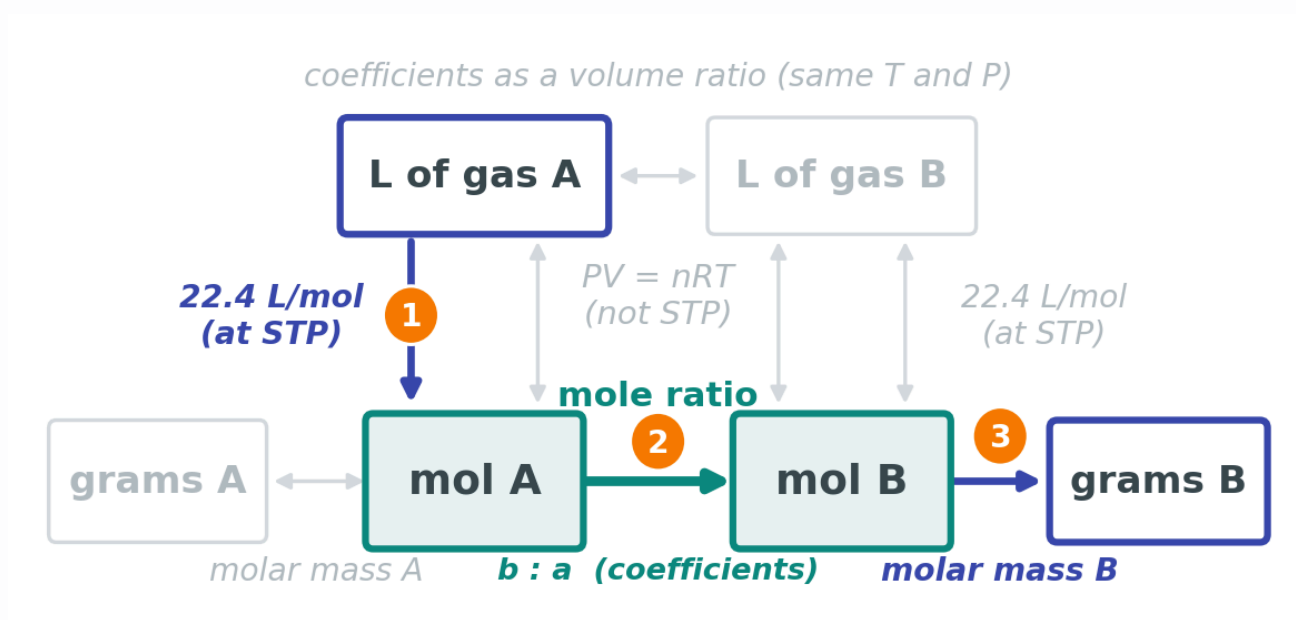
A skipped the mole ratio: $0.0300 \times 122.55 = 3.68$. B stopped at moles: $0.0300 \times (2/3) = 0.0200 \text{ mol KClO}_3$. C inverted the ratio: $0.0300 \times (3/2) \times 122.55 = 5.51$. E skipped mL → L: $672 \div 22.4 \times (2/3) \times 122.55 = 2.45 \times 10^3$.

 **Sense check:** 0.672 L is 0.0300 mol of O₂. Two KClO₃ make three O₂, so fewer moles of solid reacted: 0.0200 mol, about 2.5 g. ✓

✓ Practice 5: the route on the map



given: 672 mL O₂ = 0.672 L at STP · found: 2.45 g KClO₃



🧠 **Sense check:** mL → L comes first. Then A is O₂, the product, and B is KClO₃: molar volume, mole ratio, molar mass.

✓



Extra practice 1



molar mass $\text{C}_6\text{H}_{12}\text{O}_6$ 180.16 g/mol

Yeast in bread dough ferments glucose. **0.900 g of $\text{C}_6\text{H}_{12}\text{O}_6$** ferments completely. What volume of CO_2 , in mL, forms at STP?

- A. 112
- B. 0.224
- C. 224
- D. 56.0
- E. 0.00999

✓ Extra practice 1 answer: C



given: 0.900 g $\text{C}_6\text{H}_{12}\text{O}_6$ (180.16 g/mol) · wanted: mL CO_2 at STP

$$0.900 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6 \times \frac{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{180.16 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6} \times \frac{2 \text{ mol } \text{CO}_2}{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6} = 0.00999 \text{ mol } \text{CO}_2$$

$$0.00999 \text{ mol } \text{CO}_2 \times \frac{22.4 \text{ L}}{1 \text{ mol } \text{CO}_2} \times \frac{1000 \text{ mL}}{1 \text{ L}} = \mathbf{224 \text{ mL } \text{CO}_2 \rightarrow \text{answer C}}$$

0.900 g is 0.00500 mol of glucose. A skipped the mole ratio: $0.00500 \times 22.4 \times 1000 = 112$. B stopped at liters: 0.224 L. D inverted the ratio: $0.00500 \times (1/2) \times 22.4 \times 1000 = 56.0$. E stopped at moles: 0.00999 mol CO_2 .

 **Sense check:** Two CO_2 per glucose: a hundredth of a mole of gas, 0.224 L = 224 mL. ✓



Extra practice 2



molar mass LiOH 23.95 g/mol

A spacecraft air scrubber traps exhaled CO_2 in lithium hydroxide. It absorbs **50.4 L of CO_2** , measured at STP. What mass of LiOH, in grams, does it use?

- A. 108
- B. 53.9
- C. 5.41×10^4
- D. 4.50
- E. 26.9

✓ Extra practice 2 answer: A



given: 50.4 L CO₂ at STP · wanted: g LiOH (23.95 g/mol)

$$50.4 \text{ L CO}_2 \times \frac{1 \text{ mol CO}_2}{22.4 \text{ L CO}_2} \times \frac{2 \text{ mol LiOH}}{1 \text{ mol CO}_2} \times \frac{23.95 \text{ g LiOH}}{1 \text{ mol LiOH}} = 108 \text{ g} \rightarrow \text{answer A}$$

B skipped the mole ratio: $2.25 \times 23.95 = 53.9$. C flipped the molar volume: $50.4 \times 22.4 \times 2 \times 23.95 = 5.41 \times 10^4$; liters on top do not cancel. D stopped at moles: $2.25 \times 2 = 4.50$ mol LiOH. E inverted the ratio: $2.25 \times (1/2) \times 23.95 = 26.9$.

🧠 **Sense check:** 50.4 L at STP is 2.25 mol of CO₂. Each CO₂ takes two LiOH, so 4.50 mol of LiOH, a little over 100 g. ✓

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?

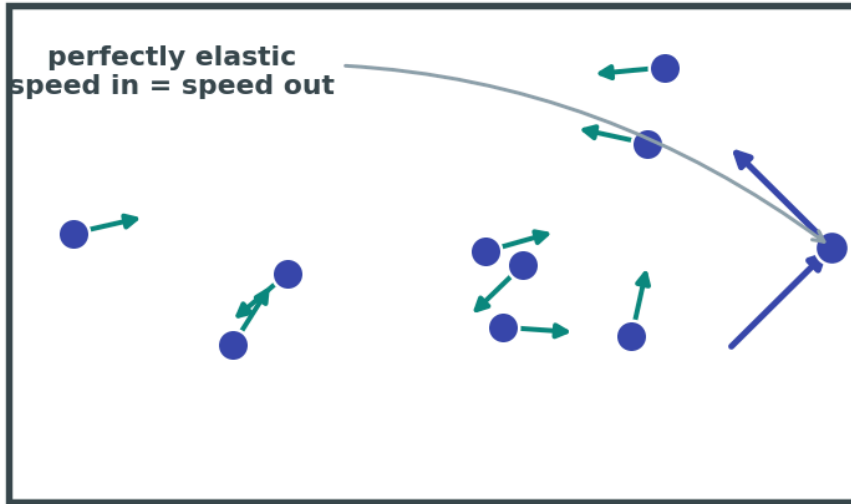
One mole of any gas fills the same 22.4 L at STP because gas particles are tiny, far apart, and moving with an average energy set by the temperature alone. The **kinetic-molecular theory** explains this behavior.

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



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.

Energy of motion: $KE = \frac{1}{2}mv^2$

$$KE = \frac{1}{2} m v^2$$

m, mass in kg · v, speed in m/s · KE in joules

A 2.0 kg cart and an 8.0 kg cart each carry 100 J. Which moves faster?

$$\frac{1}{2} \times 2.0 \text{ kg} \times (10. \text{ m/s})^2 = \mathbf{100 \text{ J}} \cdot \frac{1}{2} \times 8.0 \text{ kg} \times (5.0 \text{ m/s})^2 = \mathbf{100 \text{ J}}$$

The light cart: four times the mass, half the speed. Gas particles follow the same rule.

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 = $\sqrt{(M_{\text{heavy}} / M_{\text{light}})}$

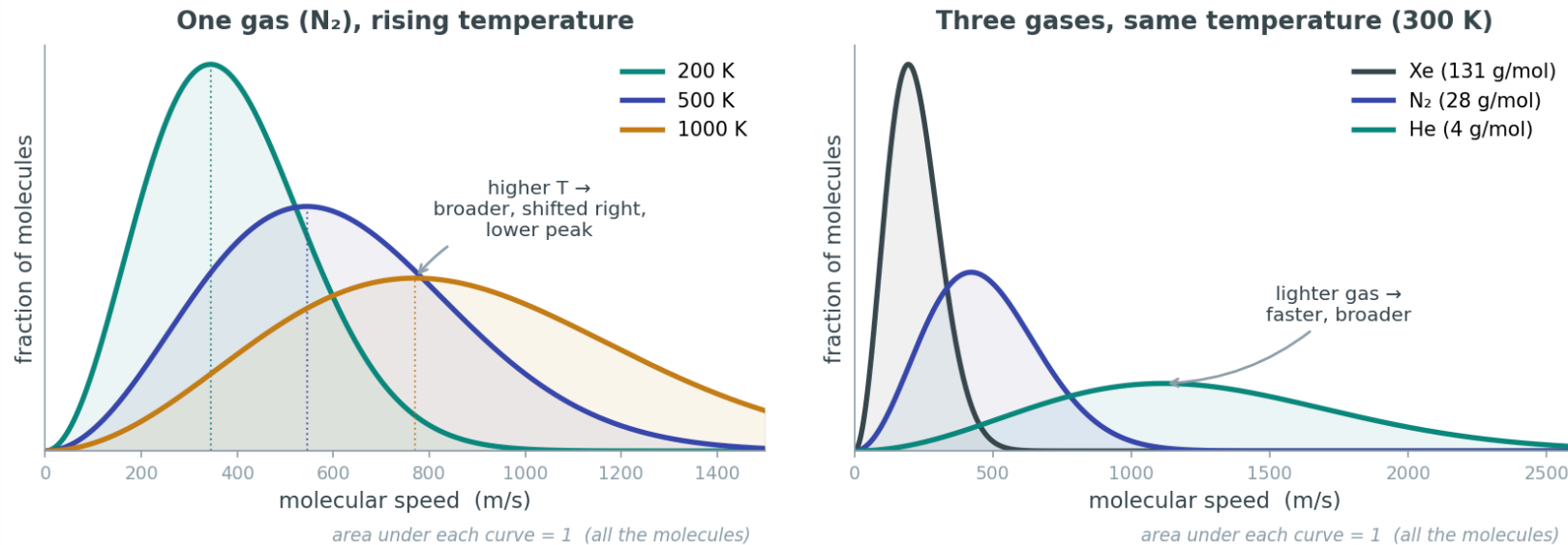
Kinetic energy is $\frac{1}{2}mu^2$. Two gases with equal average kinetic energy but different masses cannot share a speed: the lighter one moves faster.

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

The Maxwell-Boltzmann distribution

The Maxwell-Boltzmann speed distribution

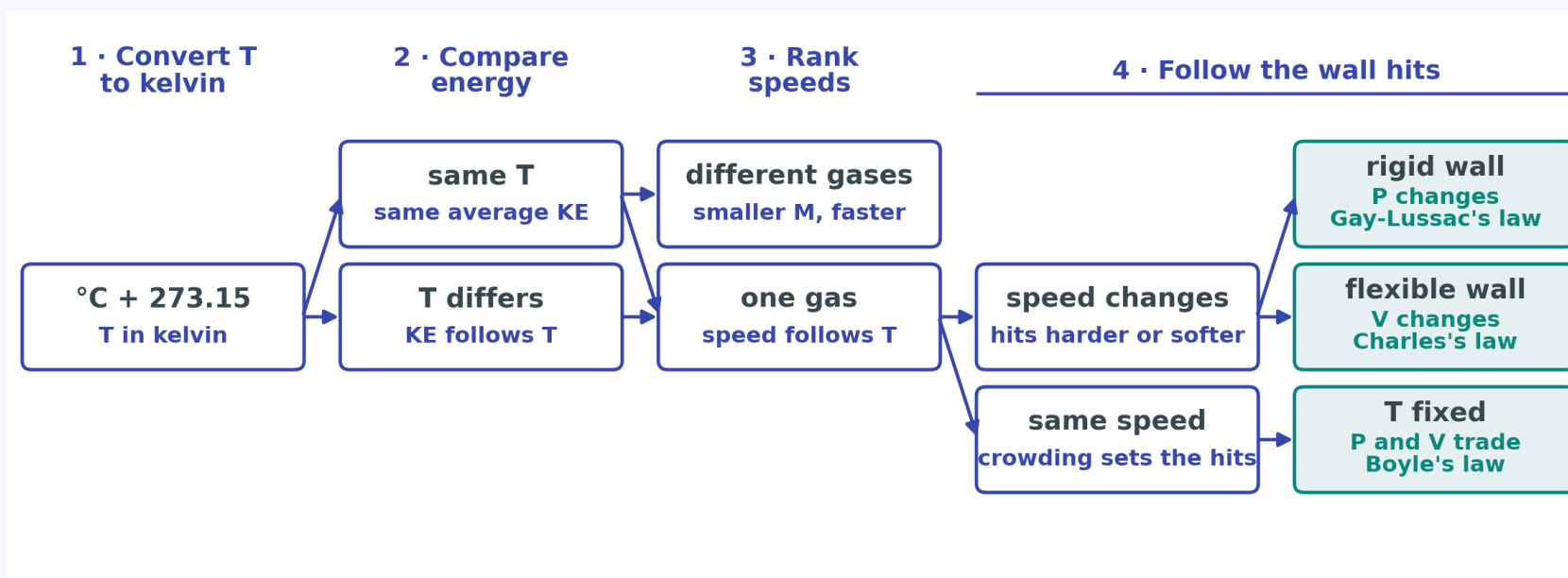


Molecules spread across a range of speeds. Heating shifts the curve right and flattens it; a heavier gas peaks at lower speed, taller and narrower. Every curve encloses the same area: all the molecules.



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 $\sqrt{T}$.
4. **Follow the wall hits.**





Worked example 1: warming one gas

$$KE_{\text{avg}} \propto T \text{ in kelvin}$$

given: a sealed steel tank of N_2 · heated from 27°C to 100°C · wanted: the change in average KE, average speed, and pressure

A sealed steel tank of nitrogen is heated from **27°C** to **100°C** . Does the average kinetic energy of the N_2 molecules rise, fall, or stay the same? What happens to their average speed, and to the pressure in the tank?

✓ Worked example 1: solution

$$KE_{\text{avg}} \propto T \text{ in kelvin}$$

one sealed steel tank of N_2 · $27^\circ\text{C} \rightarrow 100^\circ\text{C}$ · the volume cannot change

Step 1 · Convert T to kelvin

$$27 + 273.15 = 300.15 \text{ K} \rightarrow 100 + 273.15 = 373.15 \text{ K}$$

Step 2 · Compare energy

$$300.15 \text{ K} \rightarrow 373.15 \text{ K} \Rightarrow \text{average KE rises (KE follows T)}$$

Step 3 · Rank speeds

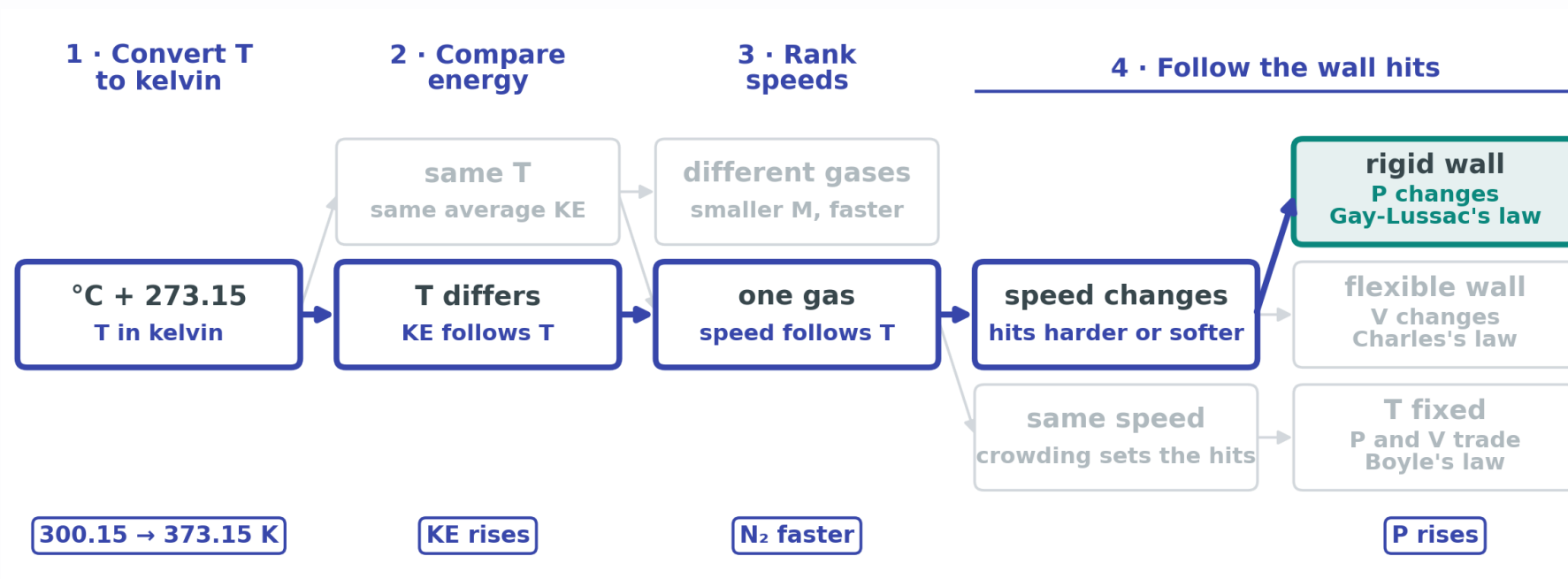
same N_2 , same mass, more KE $\Rightarrow$ average speed rises

🧠 **Sense check:** **Step 4 · Follow the wall hits** Harder, more frequent hits on rigid walls: P rises (Gay-Lussac). ✓

✓ Worked example 1: the route on the strip

$$KE_{avg} \propto T \text{ in kelvin}$$

given: a sealed steel tank of N_2 · 27 °C → 100 °C · found: KE rises · speed rises · P rises



🧠 **Sense check:** Only T changed: more KE, more speed, harder hits, higher P in a rigid tank. ✓



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

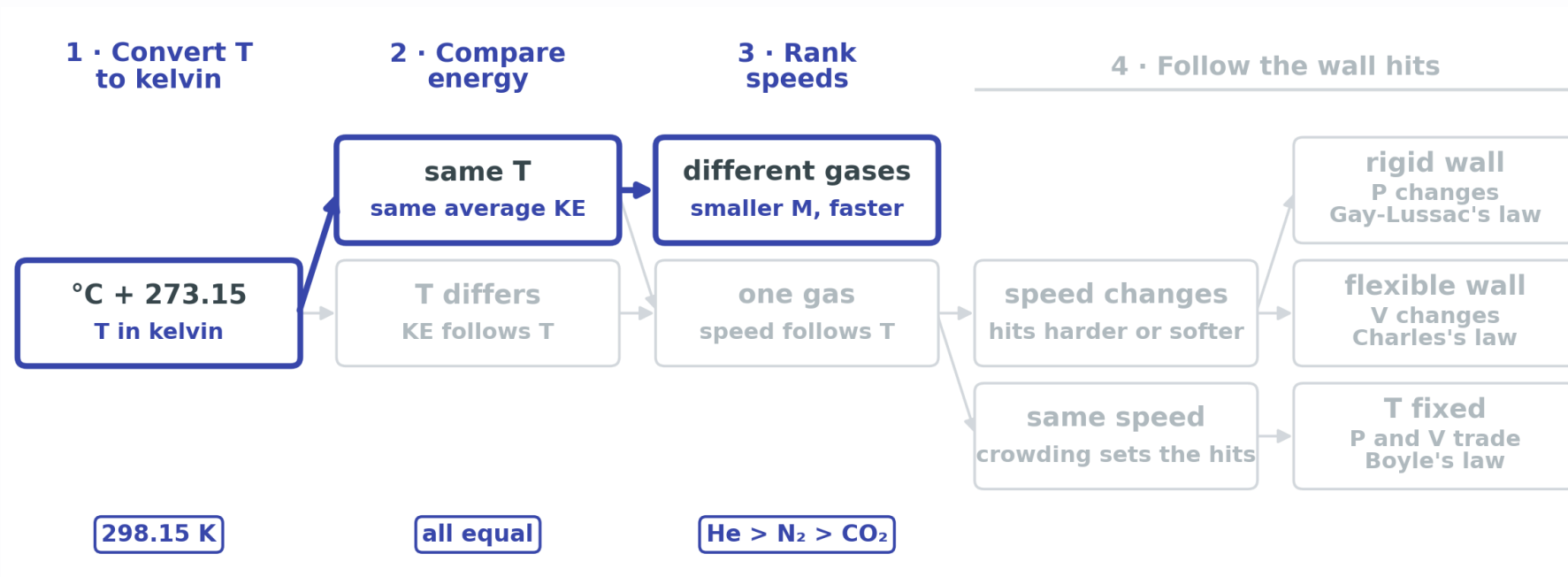
$$u(\text{He}) \div u(\text{CO}_2) = \sqrt{(44.01 \div 4.003)} = 3.32$$

 **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. $KE \propto \text{Kelvin temperature}$. 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. Warming raises both the energy and the speed, but not by the same factor: the speed grows more slowly than the energy. Say the speed rises, not that it doubles.



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

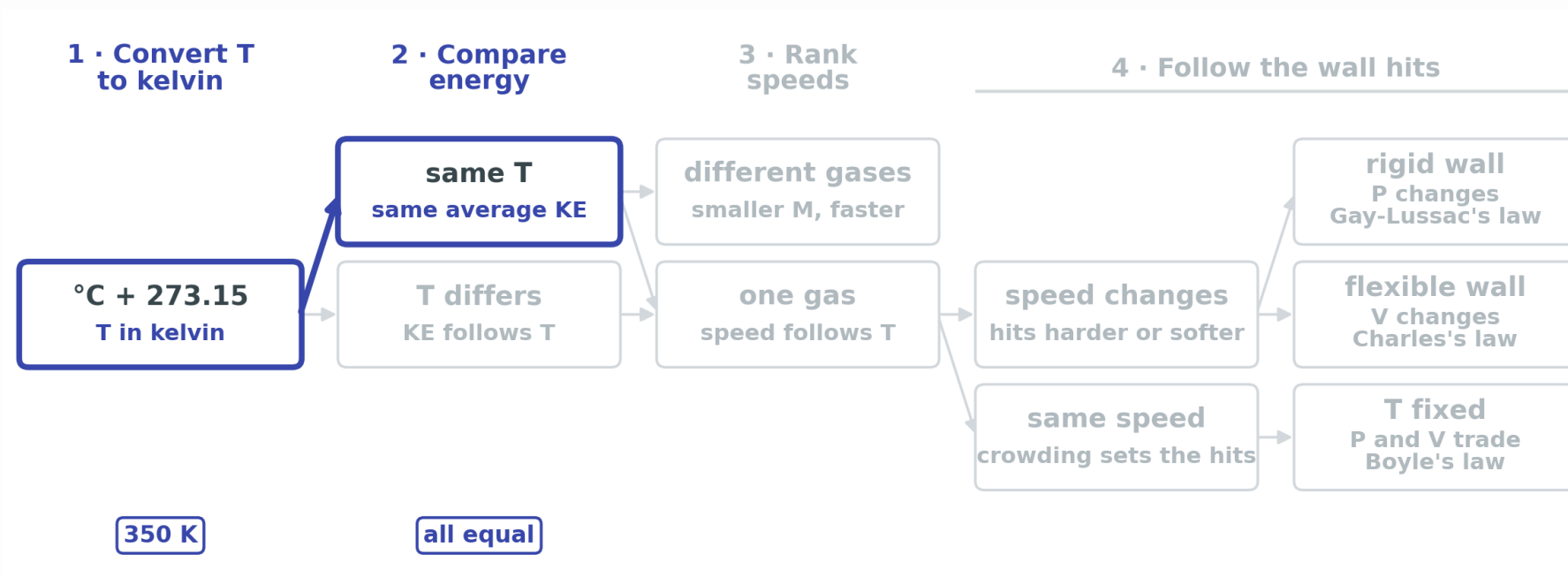
Average kinetic energy depends only on Kelvin temperature, so all three are equal. A confuses speed with energy: He is the fastest, but that is how it carries the same KE with less mass. C assumes mass sets the energy; mass sets the speed. D adds a variable that never appears: $(3/2)kT$ has no pressure 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 1: the route on the strip

$$KE_{\text{avg}} \propto T \text{ in kelvin}$$

given: He, O₂, SO₂ · all at 350 K · found: equal average KE per molecule



Sense check: The question asked about energy only. One shared temperature settles it at Step 2. ✓



Practice 2: a balloon in the freezer

$$KE_{\text{avg}} \propto T \text{ in kelvin}$$

given: a sealed party balloon · 21 °C room → −18 °C freezer · the outside air pressure is unchanged

A sealed party balloon is moved from a **21 °C** room into a **−18 °C** freezer. What happens to the average speed of its gas particles, and to its volume?

- A. Speed falls; volume shrinks
- B. Speed falls; volume stays the same
- C. Speed stays the same; volume shrinks
- D. Speed stays the same; volume stays the same

✓ Practice 2 answer: A

$$21 + 273.15 = 294.15 \text{ K} \rightarrow -18 + 273.15 = 255.15 \text{ K}$$

lower T, less average KE · same gas: slower particles · softer, fewer hits · the flexible balloon shrinks →
answer A

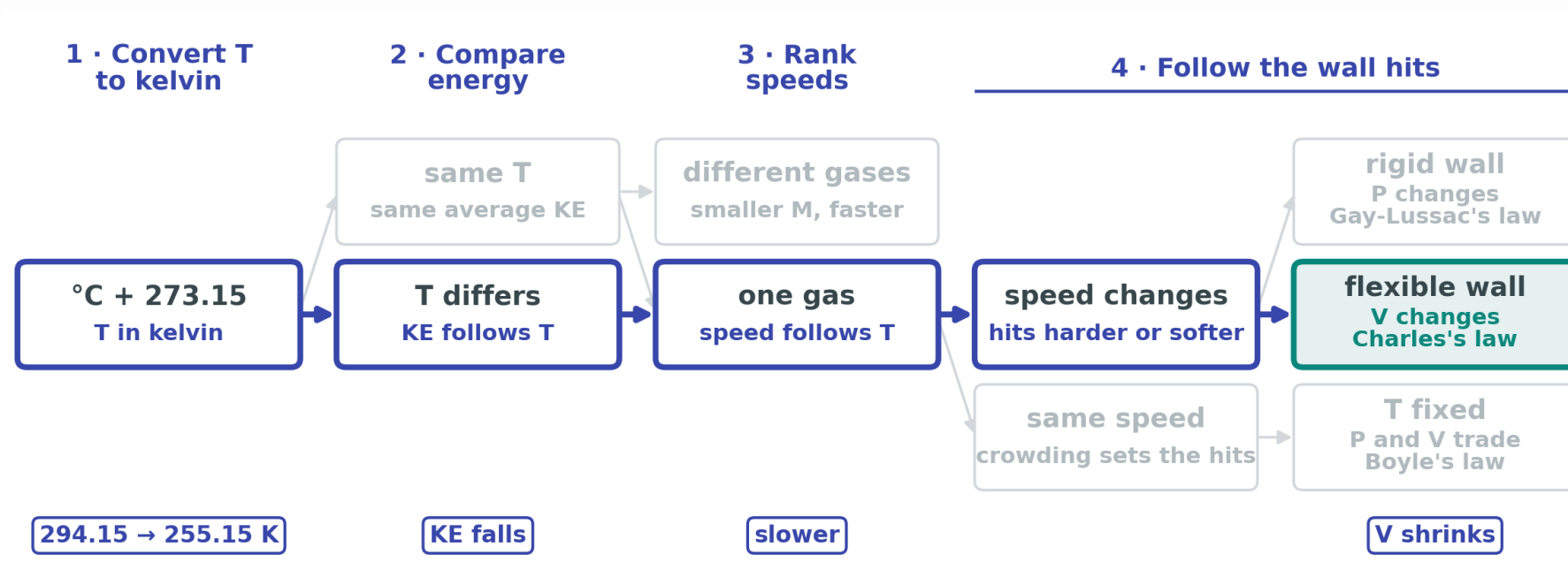
B treated the balloon as a rigid tank. Its wall is flexible, so it shrinks until the inside hits balance the outside air again. C kept the speed fixed, but KE follows T: colder particles move slower. The particles themselves never shrink. D held everything fixed because no gas left. The particles stay, but they slow down.

🧠 **Sense check:** Charles's law puts a number on it: the volume falls to $255.15 \div 294.15 = 0.867$ of its start. ✓

✓ Practice 2: the route on the strip

$$KE_{avg} \propto T \text{ in kelvin}$$

given: a sealed balloon · 21 °C → −18 °C · found: speed falls · volume shrinks



🧠 **Sense check:** Worked example 1's path, run cold, ends at a flexible wall: V changes, not P. ✓

 Practice 3

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

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 3 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

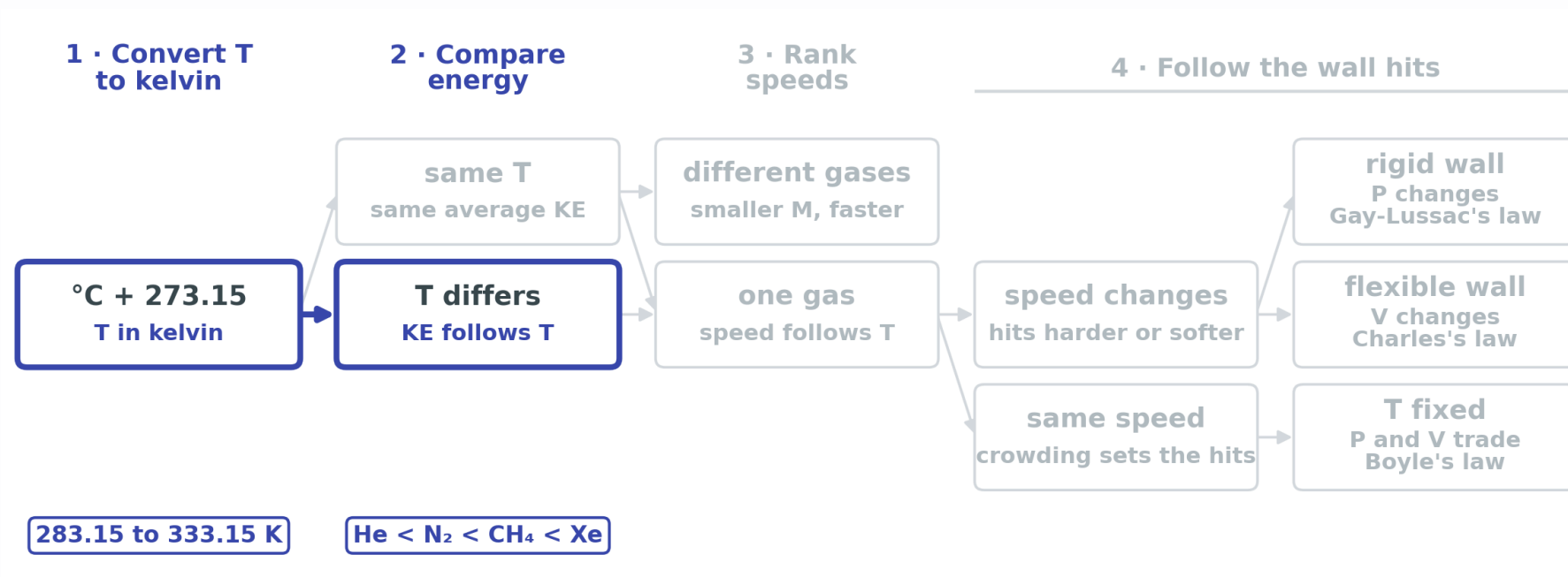
A applied the same-temperature rule, but these flasks sit at four different temperatures. B compared the raw numbers, setting 300 K against Celsius readings; convert every T to kelvin first. C ranked by speed, lightest fastest, but speed is not energy: mass never enters $(3/2)RT$.

 **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: the route on the strip

$$KE_{avg} \propto T \text{ in kelvin}$$

given: He 10 °C · N₂ 300 K · CH₄ 30 °C · Xe 60 °C · found: He < N₂ < CH₄ < Xe



🧠 **Sense check:** Four temperatures, so KE follows T. The gases' masses never enter an energy ranking. ✓



Practice 4

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

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 4 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

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

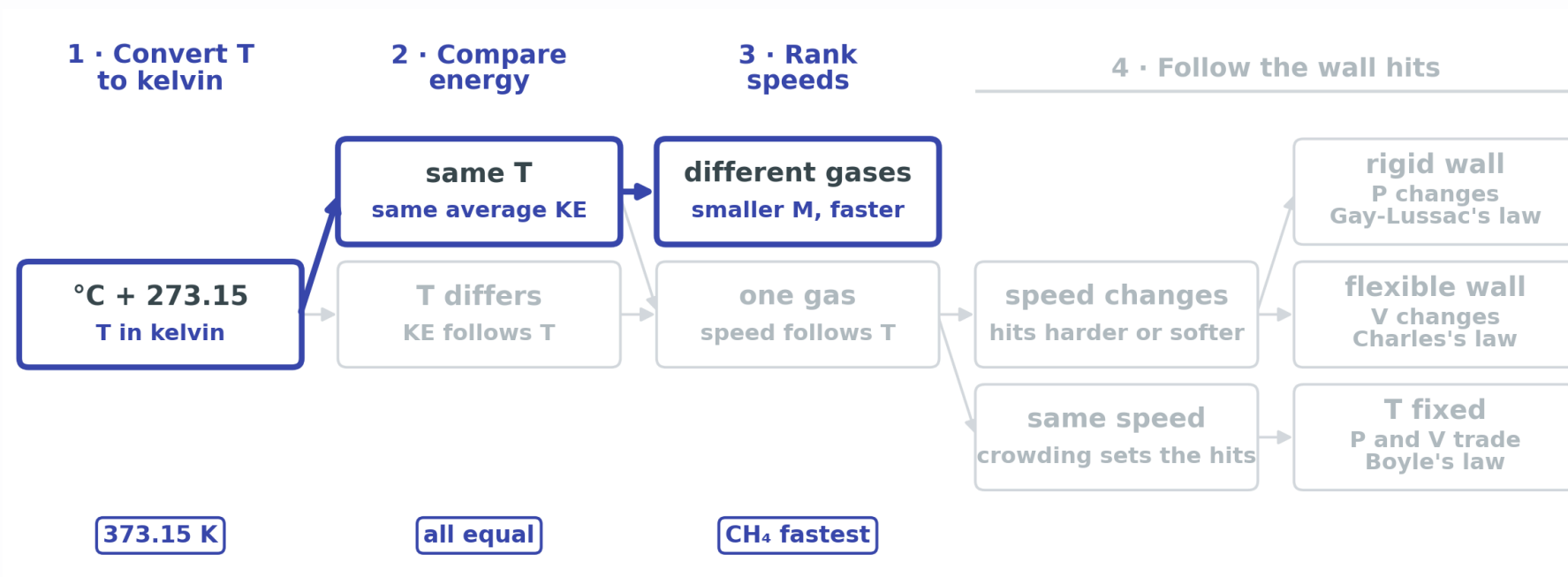
A picked the heaviest gas, which is the slowest. B picked the single atom, but Ar (39.95 g/mol) is heavier than the two light molecules. D is close, but 17.03 > 16.04, so ammonia trails methane by 3%. E is true of the average kinetic energy, not the speed.

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

✓ Practice 4: the route on the strip

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

given: CH_4 , NH_3 , Ar , SO_2 · all at 100°C · found: CH_4 fastest



🧠 **Sense check:** The path of worked example 2: one temperature, so the smallest molar mass is the fastest. ✓



Worked example 3: squeezing a syringe

pressure = hits per second $\times$ force per hit

given: a sealed syringe of air $\cdot$ 60.0 mL $\rightarrow$ 20.0 mL $\cdot$ held at 20 °C $\cdot$ wanted: the change in average KE and speed, and why P rises

The tip of a syringe holding **60.0 mL** of air is sealed, and the plunger is pushed in to **20.0 mL**. The syringe stays at **20 °C**. Explain the pressure rise with the model.

A common first answer: squeezing speeds the particles up. Test it.

✓ Worked example 3: energy and speed

pressure = hits per second $\times$ force per hit
a sealed syringe of air $\cdot$ 60.0 mL $\rightarrow$ 20.0 mL $\cdot$ held at 20 °C

Step 1 $\cdot$ Convert T to kelvin

$20 + 273.15 = \mathbf{293.15\text{ K}}$ before and after the squeeze

Step 2 $\cdot$ Compare energy

$293.15\text{ K} \rightarrow 293.15\text{ K} \Rightarrow$ same average KE

Step 3 $\cdot$ Rank speeds

same air, same T $\Rightarrow$ same average speed

 **Sense check:** The squeeze does not speed the particles up. Only a temperature change could. ✓

✓ Worked example 3: the wall hits

pressure = hits per second $\times$ force per hit

a sealed syringe of air $\cdot$ 60.0 mL $\rightarrow$ 20.0 mL $\cdot$ 293.15 K throughout $\cdot$ same average speed

Step 4 $\cdot$ Follow the wall hits

60.0 mL $\div$ 20.0 mL = 3: one third the room

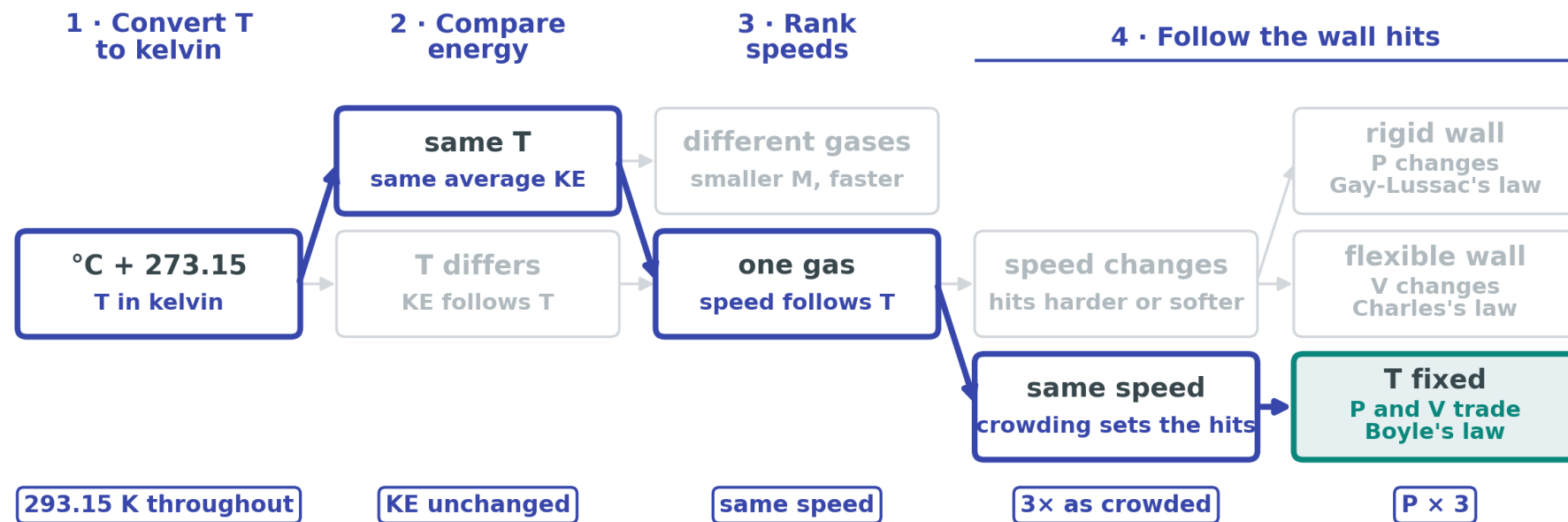
same speed, 3 times as crowded $\cdot$ 3 times the hits on each bit of wall $\cdot$ each hit no harder $\cdot$ $P \times 3$

 **Sense check:** The hits came more often, not harder. The pressure rose because the room shrank at fixed T: Boyle's law. ✓

✓ Worked example 3: the route on the strip

pressure = hits per second $\times$ force per hit

given: 60.0 mL $\rightarrow$ 20.0 mL $\cdot$ held at 20 °C $\cdot$ found: same KE $\cdot$ same speed $\cdot$ P $\times$ 3



🧠 **Sense check:** Same temperature, so Steps 2 and 3 change nothing. Crowding alone raised the pressure. ✓

Take-home: only temperature changes the speed

squeeze it · add more gas · let it expand at fixed $T \Rightarrow$ same average speed
heat it or cool it $\Rightarrow$ the speed changes · the hits follow both speed and crowding

Pushing on a gas changes how crowded its particles are, not how fast they move. Only a change in temperature changes the average speed.

Practice 5: a bag of chips on a mountain road

pressure = hits per second $\times$ force per hit

given: a sealed bag of chips · held at 15 °C · the outside air pressure falls on the climb

A sealed bag of chips puffs up as a car climbs a mountain road, though the bag stays at **15 °C**. Why does it puff up, and which gas law describes the change?

- A. Inside particles speed up and hit harder · Charles's law
- B. Inside particles swell as the pressure drops · Boyle's law
- C. More particles form inside the bag · Avogadro's law
- D. Fewer outside hits; same inside speed · Gay-Lussac's law
- E. Fewer outside hits; same inside speed · Boyle's law

✓ Practice 5 answer: E

$15 + 273.15 = 288.15 \text{ K}$ the whole climb $\Rightarrow$ same average KE, same inside speed
thinner air outside: fewer hits on the outside of the bag · the bag grows until the inside hits thin to
match · T fixed, V up as P falls: Boyle's law $\rightarrow$ answer E

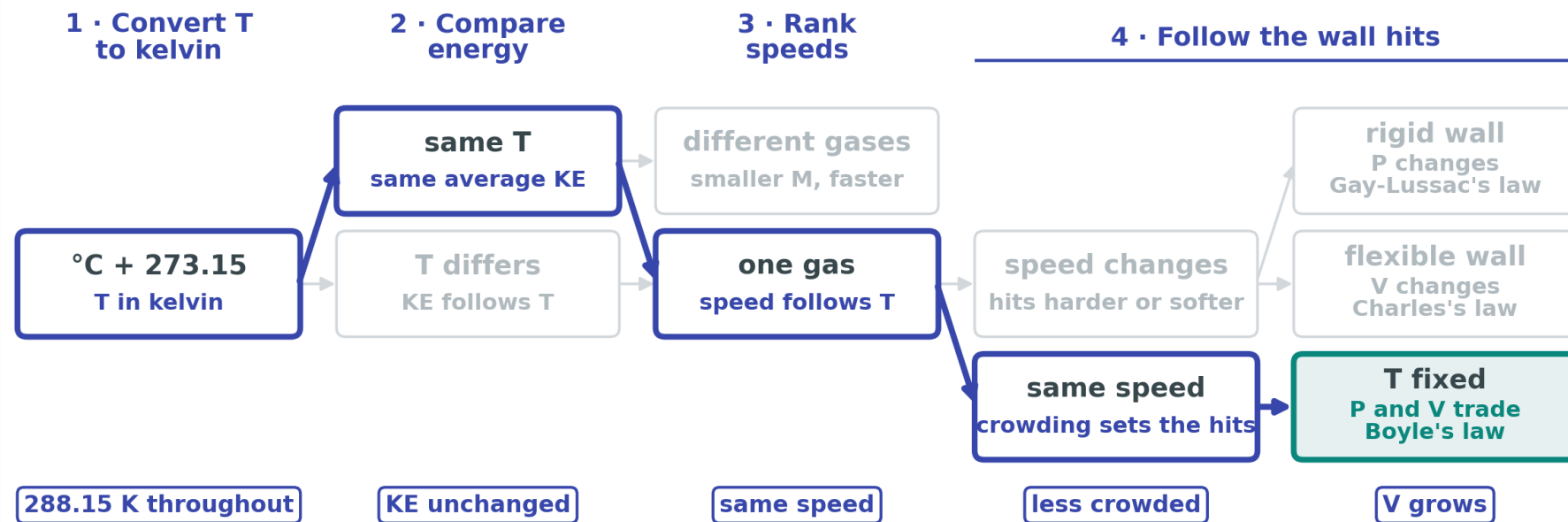
A blamed a temperature change, but the bag stays at 15°C , so the speed holds. B made the particles swell; their size is fixed and negligible. C added particles to a sealed bag. D has the right reason and the wrong law: Gay-Lussac's law needs a fixed volume, and this volume grew.

 **Sense check:** Same speed, more room: the inside gas spreads out until its hits match the thinner air outside. ✓

✓ Practice 5: the route on the strip

pressure = hits per second $\times$ force per hit

given: a sealed bag $\cdot$ 15 °C throughout $\cdot$ found: same speed $\cdot$ V grows $\cdot$ Boyle's law



🧠 **Sense check:** Worked example 3's path, run the other way: less crowding, a larger volume. ✓



Practice 6

$$P_{\text{total}} = P_{\text{O}_2} + P_{\text{Kr}}$$

given: 0.150 mol O₂ + 0.150 mol Kr · one rigid flask · 35 °C · wanted: partial pressures, average KE, average speeds

A rigid flask holds **0.150 mol of oxygen gas (O₂)** and **0.150 mol of krypton** at **35 °C**. Which comparison of the two gases is correct?

- A. $P_{\text{O}_2} > P_{\text{Kr}}$ · equal average KE · O₂ faster
- B. $P_{\text{O}_2} = P_{\text{Kr}}$ · equal average KE · O₂ faster
- C. $P_{\text{Kr}} > P_{\text{O}_2}$ · equal average KE · O₂ faster
- D. $P_{\text{O}_2} = P_{\text{Kr}}$ · equal average KE · equal speeds
- E. $P_{\text{O}_2} = P_{\text{Kr}}$ · O₂ more average KE · O₂ faster

✓ Practice 6 answer: B

$P_{\text{total}} = P_{\text{O}_2} + P_{\text{Kr}}$ · each gas acts as if it were alone · $\text{KE}_{\text{avg}} \propto T$ in kelvin

35 + 273.15 = 308.15 K for both · same 0.150 mol, same flask $\Rightarrow P_{\text{O}_2} = P_{\text{Kr}}$ · same T $\Rightarrow$ same average KE
 · O_2 32.00 (2×16.00) < Kr 83.80 g/mol $\Rightarrow \text{O}_2$ faster $\rightarrow$ answer B

A counted only how often the particles hit: the lighter O_2 molecules hit more often, but each hit is softer. C counted only how hard they hit: krypton atoms hit harder, but less often. The two effects cancel. D gave equal energies equal speeds; heavier krypton moves slower to carry the same KE. E gave the faster gas more energy; at one temperature every gas has the same average KE.

 **Sense check:** A partial pressure depends on moles, T and V, never on which gas it is. ✓



Practice 7

	statement
1	Two gas particles that collide keep their total kinetic energy
2	Gas particles attract one another, so each slows just before it hits a wall
3	At 500 K, SF_6 molecules and H_2 molecules have the same average kinetic energy

Three statements about gas particles. Which statements agree with kinetic-molecular theory?

- A. Statement 1 only
- B. Statement 3 only
- C. Statements 1, 2 and 3
- D. Statements 1 and 3 only
- E. Statements 2 and 3 only

✓ Practice 7 answer: D

agree: 1 and 3 · contradict: 2 → answer D

1: collisions are perfectly elastic ✓ · 2: no forces act between particles ✗ · 3: average KE follows Kelvin T alone, 500 K for both ✓

A rejected statement 3, giving the heavier SF_6 more energy; at one temperature every gas has the same average KE, and SF_6 simply moves slower. B rejected statement 1, as if collisions used up energy; elastic collisions never let the gas run down. C accepted statement 2; the model puts no forces between particles. E accepted statement 2 and rejected statement 1.

🧠 **Sense check:** Statements 1 and 3 restate two of the five postulates. Statement 2 breaks a third. ✓



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. Does its average kinetic energy double? Does its average speed double too?

The five postulates are the microscopic story behind every gas law.

Can you...?

- ☐ describe gas behavior with kinetic-molecular theory and convert pressure among atm, mmHg, kPa, and psi?
- ☐ apply Boyle's, Charles's, Gay-Lussac's, Avogadro's, and the combined gas law to a fixed gas sample, with temperature in kelvin?
- ☐ use $PV = nRT$ to find pressure, volume, amount, or temperature, and the density or molar mass of a gas?
- ☐ convert between moles and liters of a gas at STP with 22.4 L/mol and carry gas volumes through stoichiometry?
- ☐ apply Dalton's law to gas mixtures, including a gas collected over water?
- ☐ compare gases at one temperature: same average kinetic energy, and the lighter gas moves, diffuses, and effuses faster?

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