

The Mole

Preparation for General Chemistry · Dr. Karmach

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

- Calculate the molar mass of a compound from its formula
- Convert among grams, moles, and particles with molar mass and Avogadro's number, including atoms or ions of one element inside a compound
- Calculate the percent composition of a compound from its formula or from measured masses, and use it to find the mass of an element in a sample
- Reduce a molecular formula to its empirical formula, and find an empirical formula from percent or mass data
- Scale an empirical formula to the molecular formula with the molar mass

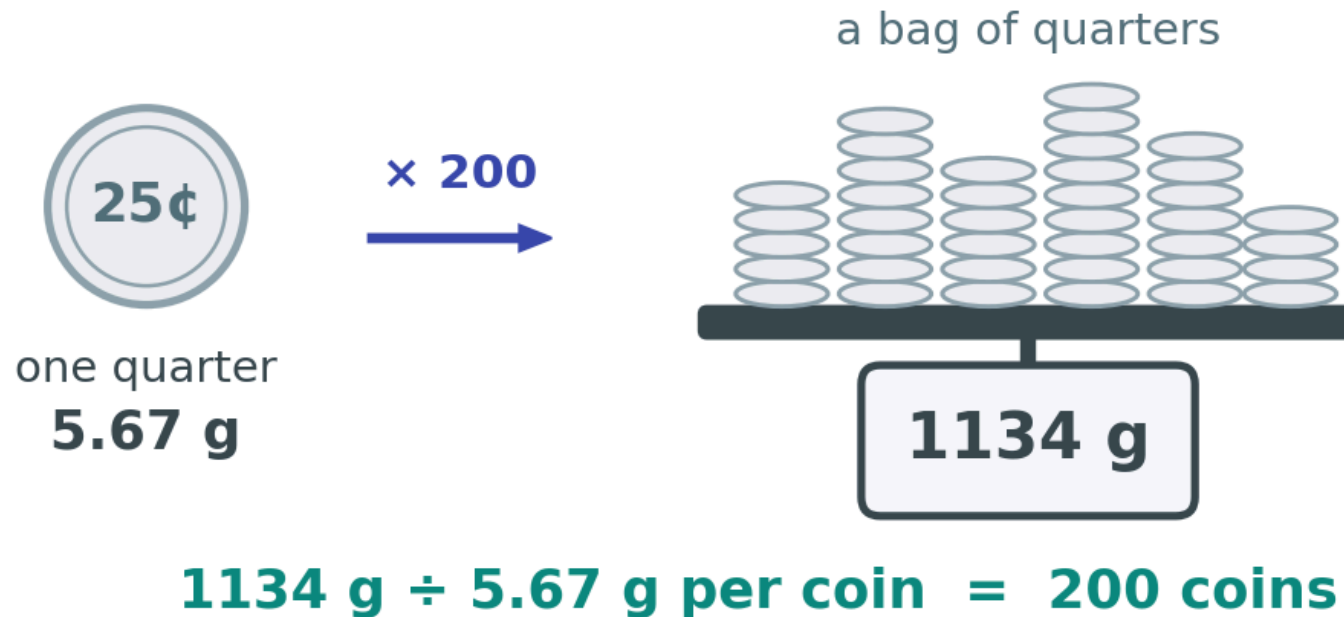
Today's route

1. **Grams, Moles & Particles**
2. **Percent Composition**
3. **Empirical & Molecular Formulas**
4. **Percent Concentration**
5. **Molarity**
6. **Dilution**

1 • Grams, Moles & Particles

Convert among grams, moles, particle counts, and atoms of one element, using the molar mass built from the formula, Avogadro's number, and the formula's subscripts.

Counting by weighing



A bank counts coins by weighing them. One quarter weighs **5.67 g**, so a 1134 g bag holds **200 quarters**. The scale reads a mass; the teller reports a count.

Chemists count atoms by weighing

one molecule of water: 2.99×10^{-23} g
far below what any balance reads

Atoms are far too small to count or weigh one at a time. Chemistry counts them in fixed-size batches: weigh the sample, and the mass gives the count.

The mole

$$1 \text{ mol} = 6.022 \times 10^{23} \text{ particles}$$

Avogadro's number: atoms, molecules, or formula units; the formula names the particle

The mole is chemistry's counting unit. A dozen is 12 of anything; a mole is 6.022×10^{23} of anything: a batch big enough to weigh.

Moles and particles: a counting factor

$$1 \text{ dozen} = 12 \cdot 1 \text{ ream} = 500 \cdot 1 \text{ mol} = 6.022 \times 10^{23}$$

each equality counts items in fixed-size batches

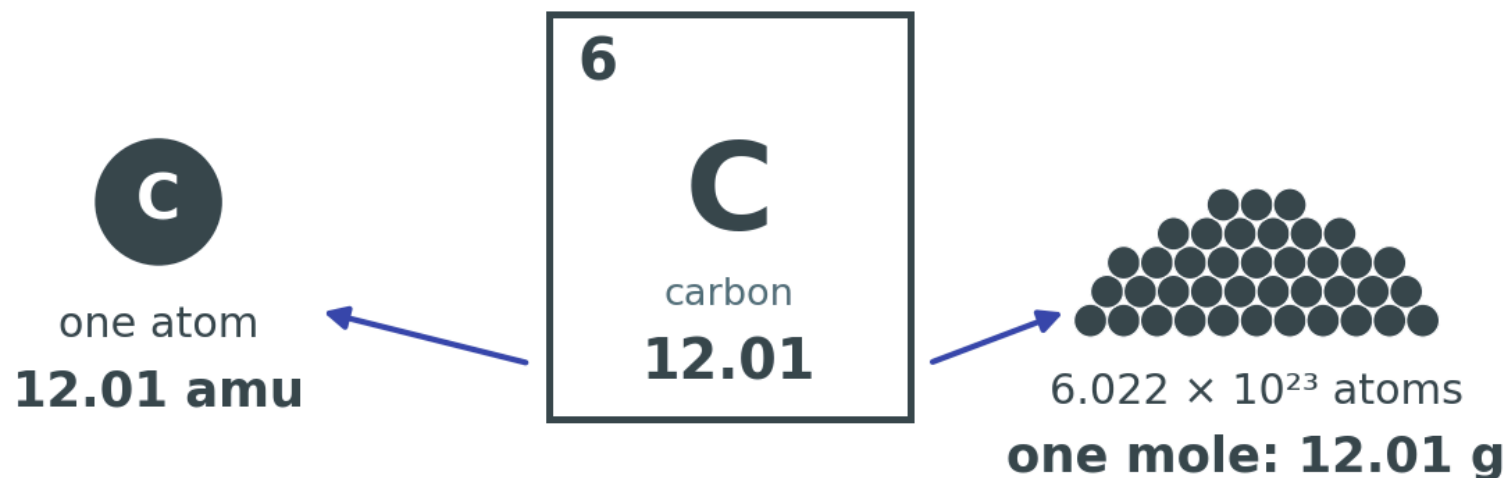
Like a dozen, a mole converts a batch count into an item count. Gold is made of one kind of particle, the gold atom:

$$3.00 \text{ mol Au} \times \frac{6.022 \times 10^{23} \text{ Au atoms}}{1 \text{ mol Au}} = 1.81 \times 10^{24} \text{ Au atoms}$$

How many moles are 3.01×10^{23} water molecules?

$$3.01 \times 10^{23} \text{ molecules} \times \frac{1 \text{ mol H}_2\text{O}}{6.022 \times 10^{23} \text{ molecules}} = 0.500 \text{ mol H}_2\text{O}$$

The periodic table reads in grams per mole



the same number reads twice: amu for one atom, grams for one mole

Avogadro's number is sized so one atom's mass in amu equals one mole's mass in grams. Every atomic mass on the periodic table is also a molar mass, in g/mol.

Molar mass of a compound

A compound's molar mass adds every atom in the formula. Subscripts multiply; a subscript outside parentheses multiplies everything inside.

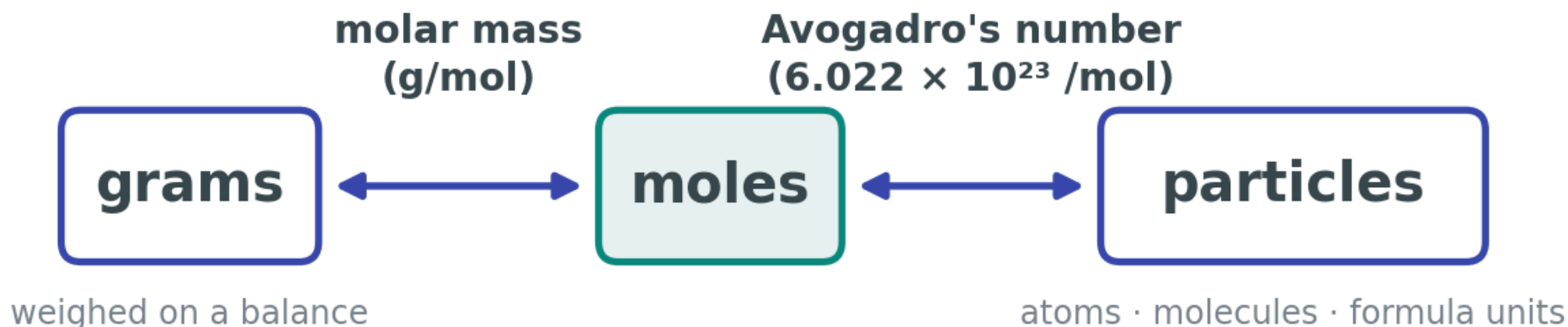
$$\text{H}_2\text{O}: 2(1.008) + 16.00 = 18.02 \text{ g/mol}$$

2 H + 1 O · subscripts count atoms

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

1 C + 2 O

One route: grams, moles, particles



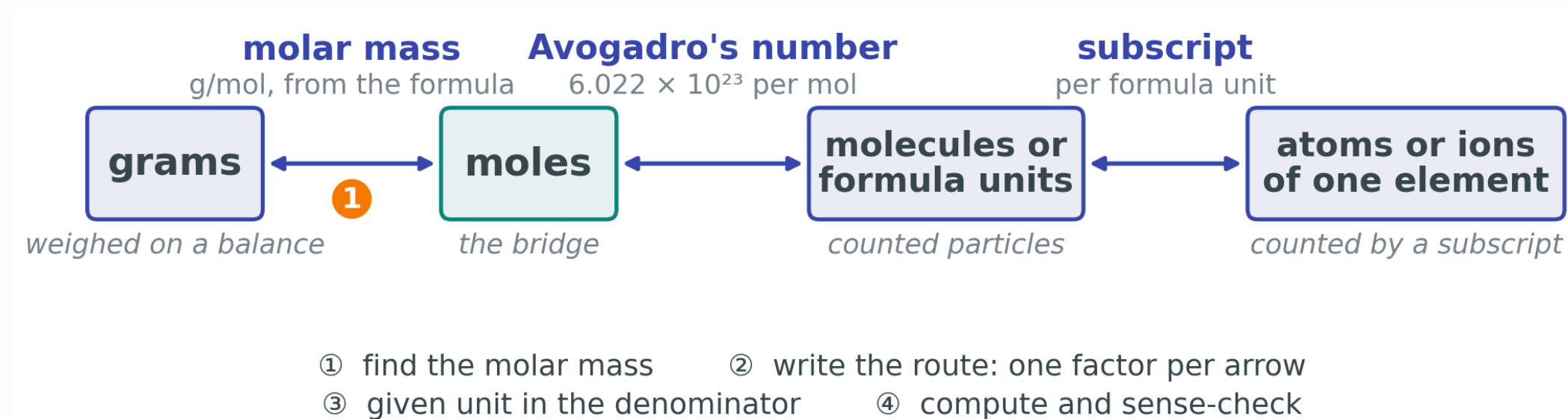
Molar mass links grams to moles. Avogadro's number links moles to particles. Every conversion between a mass and a particle count runs through moles.



The method

1. **Find the molar mass:** add every atom in the formula.
2. **Write the route:** grams $\rightleftharpoons$ moles $\rightleftharpoons$ particles, one factor per arrow.
3. **Pick the orientation that cancels:** the given unit goes in the denominator.
4. **Compute and sense-check.**

The method on the map



Step 1 builds the molar mass, the factor for the first arrow. Step 2 runs from the given box to the wanted box, one factor per arrow. Every route passes through moles.



Guided example: an aluminum can

Step 1 • Find the molar mass

Al: 26.98 g/mol, read from the periodic table

given: 0.500 mol Al • wanted: g Al

An empty soda can is made of about **0.500 mol of aluminum**. What is its mass, in grams?

On the map, the route starts at moles and ends at grams.

✓ Guided example: solution

Al: 26.98 g/mol

given: 0.500 mol Al · wanted: g Al

Step 2 · Write the route

Step 3 · Pick the orientation that cancels

Step 4 · Compute and sense-check

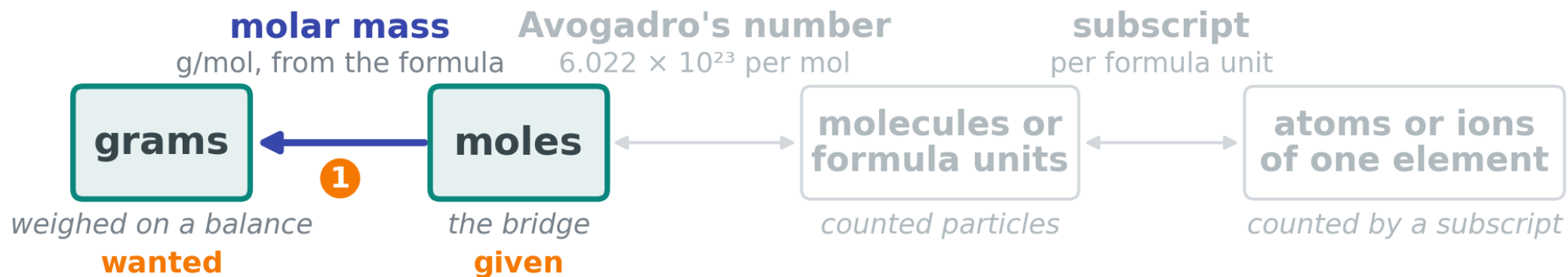
$$0.500 \cancel{\text{ mol Al}} \times \frac{26.98 \text{ g Al}}{1 \cancel{\text{ mol Al}}} = 13.5 \text{ g Al}$$

🧠 **Sense check:** Half a mole of aluminum weighs half of 26.98 g: 13.5 g, about the mass of an empty can. ✓

✓ Guided example: the route on the map

Al: 26.98 g/mol

given: 0.500 mol Al · found: 13.5 g Al



- ① find the molar mass
- ② write the route: one factor per arrow
- ③ given unit in the denominator
- ④ compute and sense-check

Sense check: The route runs right to left, across the molar-mass arrow only. Moles to grams: the molar mass goes in with grams on top. ✓



Practice 1

CaCl_2 : calcium chloride

Ca 40.08 · Cl 35.45

A closet moisture trap holds **2.50 mol of calcium chloride**. What mass of CaCl_2 is that, in grams?

- A. 0.0225
- B. 44.4
- C. 189
- D. 277

✓ Practice 1: answer D

$$\text{CaCl}_2: 40.08 + 2(35.45) = 110.98 \text{ g/mol}$$

1 Ca + 2 Cl · given: 2.50 mol CaCl₂ · wanted: g CaCl₂

$$2.50 \cancel{\text{ mol CaCl}_2} \times \frac{110.98 \text{ g CaCl}_2}{1 \cancel{\text{ mol CaCl}_2}} = \mathbf{277 \text{ g CaCl}_2 \text{ (answer D)}}$$

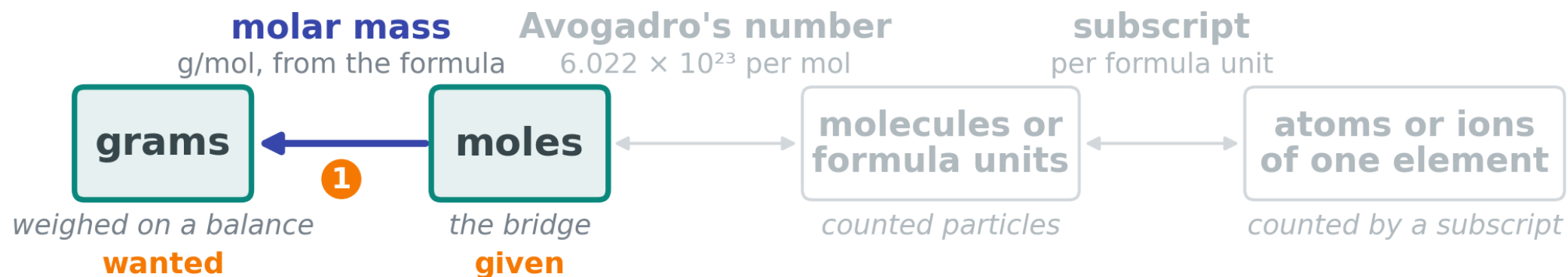
A divided by the molar mass: $2.50 \div 110.98 = 0.0225$, the grams-to-moles direction. B put the molar mass over the moles: $110.98 \div 2.50 = 44.4$. C dropped the subscript 2: $40.08 + 35.45 = 75.53 \text{ g/mol}$, and $2.50 \times 75.53 = 189$.

🧠 **Sense check:** 2.50 mol is two and a half molar masses: $2.50 \times 111 \text{ g}$ is near 280 g. ✓

✓ Practice 1: the route on the map

CaCl_2 : 110.98 g/mol

given: 2.50 mol CaCl_2 · found: 277 g CaCl_2



- ① find the molar mass
- ② write the route: one factor per arrow
- ③ given unit in the denominator
- ④ compute and sense-check

🧠 **Sense check:** One arrow, moles to grams. Step 1 counted both chlorines before the arrow was crossed. ✓



Worked example 1

$\text{Ca}(\text{NO}_3)_2$: calcium nitrate

wanted: molar mass, in g/mol · Ca 40.08 · N 14.01 · O 16.00

Calcium nitrate is a greenhouse fertilizer. Find its molar mass.

Count the atoms first: the subscript 2 sits outside the parentheses.

✓ Worked example 1: solution

$\text{Ca}(\text{NO}_3)_2$: calcium nitrate

wanted: molar mass, in g/mol · Ca 40.08 · N 14.01 · O 16.00

Step 1 · Find the molar mass

$\text{Ca}(\text{NO}_3)_2 \rightarrow 1 \text{ Ca} \cdot 2 \text{ N} \cdot 6 \text{ O}$

9 atoms per formula unit

$\text{Ca}(\text{NO}_3)_2$: $40.08 + 2(14.01) + 6(16.00) = 164.10 \text{ g/mol}$

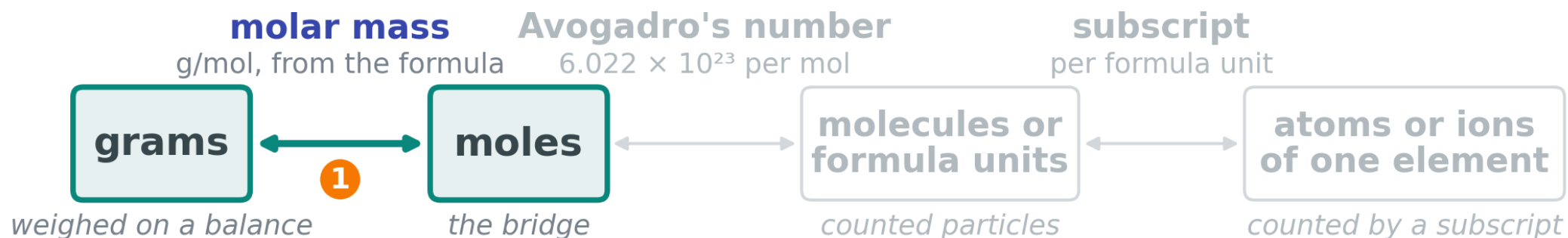
$40.08 + 28.02 + 96.00 \cdot 1 \text{ Ca} \cdot 2 \text{ N} \cdot 6 \text{ O} \cdot \text{every atom in the sum}$

 **Sense check:** Reading NO_3 once instead of twice gives 102.09, one whole NO_3 (62.01 g) short. The atom count, 1 Ca + 2 N + 6 O, catches the mistake. ✓

✓ Worked example 1: the route on the map

$\text{Ca}(\text{NO}_3)_2$: 164.10 g/mol

found: 1 mol $\text{Ca}(\text{NO}_3)_2$ = 164.10 g, the factor for the first arrow



- ① find the molar mass
- ② write the route: one factor per arrow
- ③ given unit in the denominator
- ④ compute and sense-check

Sense check: A molar mass alone is Step 1. It serves the grams-to-moles arrow in either direction, once the route is written. ✓

Worked example 2

Step 1 • Find the molar mass

$$\text{NaCl: } 22.99 + 35.45 = 58.44 \text{ g/mol}$$

given: 25.0 g NaCl • wanted: mol NaCl

A saline recipe calls for **25.0 g of table salt**. How many moles of NaCl is that?

A common first attempt uses the factor written 58.44 g over 1 mol. Test it.

✓ Worked example 2: solution

NaCl: 58.44 g/mol · given: 25.0 g NaCl · wanted: mol NaCl

A common first attempt

$$25.0 \text{ g NaCl} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = 1461 \text{ g}^2/\text{mol} \times$$

Step 2 · Write the route

Step 3 · Pick the orientation that cancels

Step 4 · Compute and sense-check

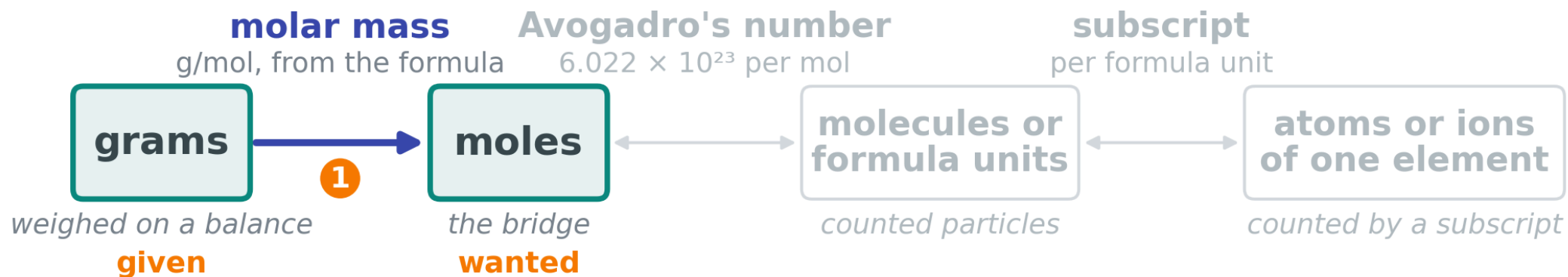
$$25.0 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 0.428 \text{ mol NaCl}$$

 **Sense check:** One mole of NaCl weighs 58.44 g. The sample weighs less than half of that, so it holds less than half a mole: 0.428. ✓

✓ Worked example 2: the route on the map

NaCl: 58.44 g/mol

given: 25.0 g NaCl · found: 0.428 mol NaCl



- ① find the molar mass
- ② write the route: one factor per arrow
- ③ given unit in the denominator
- ④ compute and sense-check

Sense check: Grams to moles crosses one arrow. The given grams sit in the denominator, so the route divides by the molar mass. ✓

Take-home: the given unit goes in the denominator

Do: grams below, so the given grams cancel.

$$25.0 \text{ g NaCl} \times (1 \text{ mol} / 58.44 \text{ g}) = 0.428 \text{ mol NaCl}$$

g cancels · mol survives ✓

Do not: grams on top. No unit cancels, and the result is not a count.

$$25.0 \text{ g NaCl} \times (58.44 \text{ g} / 1 \text{ mol}) = 1461 \text{ g}^2/\text{mol}$$

nothing cancels: the factor is upside down ✗

Your turn: carbon dioxide

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

given: 11.0 g CO₂ · wanted: mol CO₂

A soda-maker cartridge holds **11.0 g of CO₂**.

$$11.0 \text{ g CO}_2 \times \frac{\text{ } \text{mol CO}_2}{\text{ } \text{g CO}_2} = \text{ } \text{mol CO}_2$$

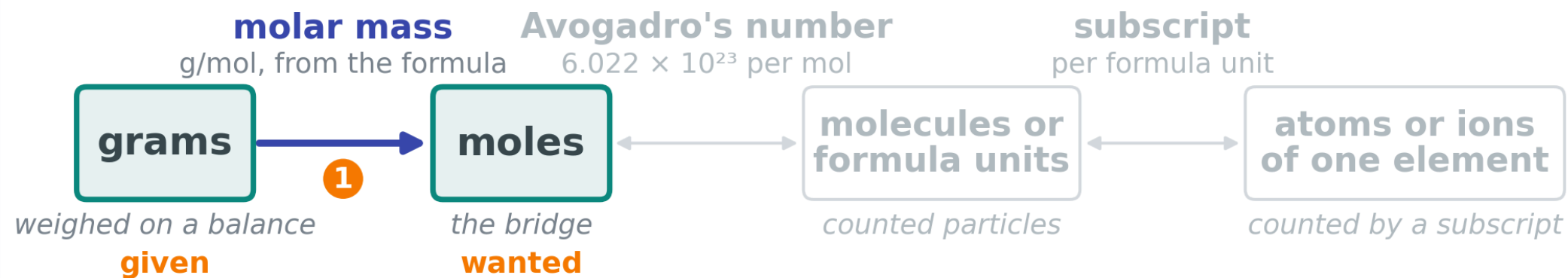
Fill the factor so the given grams cancel, then compute.

$$11.0 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} = \mathbf{0.250 \text{ mol CO}_2}$$

✓ Your turn: the route on the map

CO_2 : 44.01 g/mol

given: 11.0 g CO_2 · found: 0.250 mol CO_2



- ① find the molar mass
- ② write the route: one factor per arrow
- ③ given unit in the denominator
- ④ compute and sense-check

Sense check: One arrow, grams to moles. 11.0 g is a quarter of 44.01 g, so the cartridge holds a quarter mole. ✓

⚠ Where this goes wrong

NaCl: 58.44 g/mol

given: 25.0 g NaCl · wanted: mol NaCl

The factor upside down. $25.0 \text{ g} \times (58.44 \text{ g} / 1 \text{ mol}) = 1461 \text{ g}^2/\text{mol}$. No unit cancels, and the result is not a count. Grams must cancel: $25.0 \text{ g} \times (1 \text{ mol} / 58.44 \text{ g}) = 0.428 \text{ mol}$.

Molar mass over the sample. $58.44 \div 25.0 = 2.34$ puts the molar mass on top. A sample lighter than its molar mass holds less than one mole: 0.428, not 2.34.

Reporting particles instead of moles. Continuing with Avogadro's number, $0.428 \times 6.022 \times 10^{23} = 2.58 \times 10^{23}$, counts formula units. The question asks for moles: 0.428 mol.

 Practice 2

$$\text{MgCl}_2: 24.31 + 2(35.45) = 95.21 \text{ g/mol}$$

given: 65.0 g MgCl_2 · wanted: mol MgCl_2

Magnesium chloride de-ices winter roads. How many moles of MgCl_2 are in a **65.0 g** scoop?

- A. 0.683
- B. 1.46
- C. 6.19×10^3
- D. 65.0

✓ Practice 2: answer A

$$\text{MgCl}_2: 24.31 + 2(35.45) = 95.21 \text{ g/mol}$$

given: 65.0 g MgCl_2 · wanted: mol MgCl_2

$$65.0 \text{ g } \cancel{\text{MgCl}_2} \times \frac{1 \text{ mol MgCl}_2}{95.21 \text{ g } \cancel{\text{MgCl}_2}} = 0.683 \text{ mol MgCl}_2 \text{ (answer A)}$$

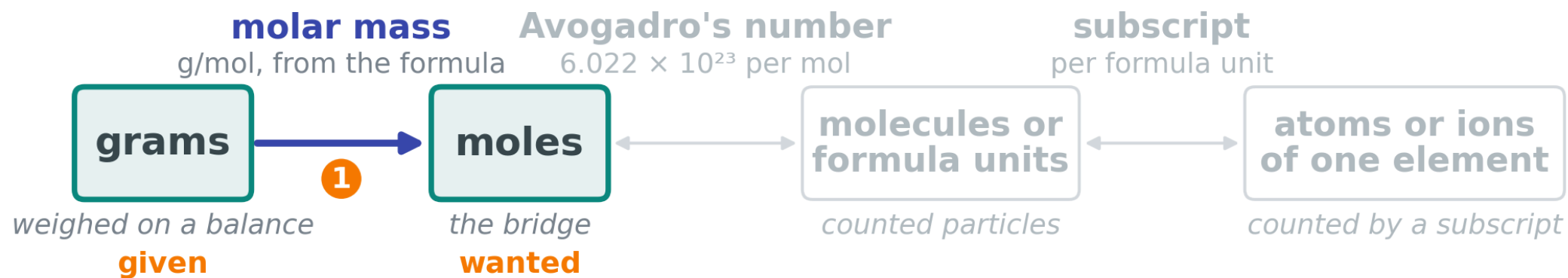
B put the molar mass on top: $95.21 / 65.0 = 1.46$. C used the factor written 95.21 g over 1 mol: $65.0 \times 95.21 = 6.19 \times 10^3$, and its real units, g^2/mol , are not moles; writing mol on it hides the flip. D skipped the molar mass: 65.0 g relabeled as 65.0 mol; grams become moles only through the g/mol factor.

🧠 **Sense check:** One mole of MgCl_2 weighs 95.21 g, and the 65.0 g scoop is about two-thirds of that: 0.683 mol. ✓

✓ Practice 2: the route on the map

MgCl_2 : 95.21 g/mol

given: 65.0 g MgCl_2 · found: 0.683 mol MgCl_2



- ① find the molar mass
- ② write the route: one factor per arrow
- ③ given unit in the denominator
- ④ compute and sense-check

🧠 **Sense check:** One arrow, grams to moles. The question stops at moles, so the route stops in the middle box. ✓



Worked example 3: grams to particles

Step 1 • Find the molar mass

$$\text{H}_2\text{O}: 2(1.008) + 16.00 = 18.02 \text{ g/mol}$$

given: 0.0500 g H₂O • wanted: molecules of H₂O

One drop of water from an eyedropper weighs about **0.0500 g**. How many H₂O molecules does the drop hold?

No single equality links grams to molecules. Set up the chain so each unit cancels the one before.

✓ Worked example 3: solution

H₂O: 18.02 g/mol

given: 0.0500 g H₂O · wanted: molecules of H₂O

Step 2 · Write the route

Step 3 · Pick the orientation that cancels

Step 4 · Compute and sense-check

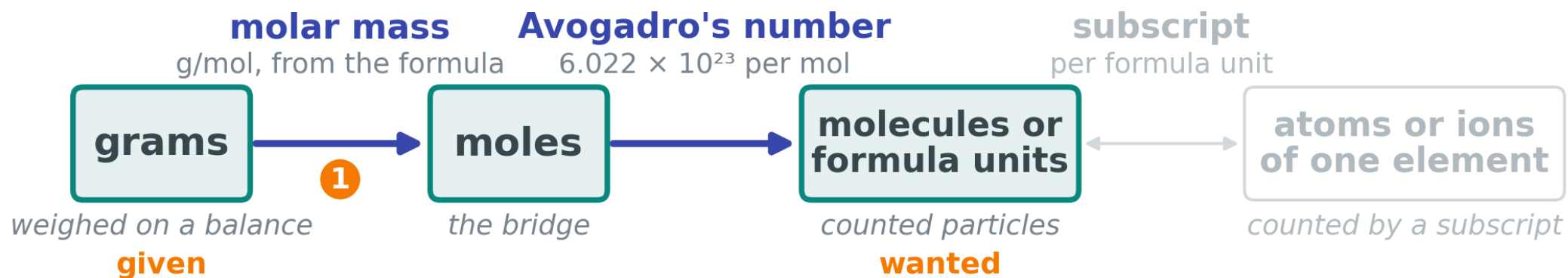
$$0.0500 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol H}_2\text{O}} = 1.67 \times 10^{21} \text{ molecules}$$

 **Sense check:** The drop is 0.00277 mol, a small fraction of a mole, so the count lands well below 6×10^{23} , yet still enormous: 1.67×10^{21} . ✓

✓ Worked example 3: the route on the map

H_2O : 18.02 g/mol

given: 0.0500 g H_2O · found: 1.67×10^{21} molecules



- ① find the molar mass
- ② write the route: one factor per arrow
- ③ given unit in the denominator
- ④ compute and sense-check

Sense check: Two arrows: molar mass, then Avogadro's number. No single factor joins grams to molecules; the route passes through moles. ✓



Practice 3

$$\text{O}_2: 2(16.00) = 32.00 \text{ g/mol}$$

given: 8.00 g O₂ · wanted: molecules of O₂

A party balloon holds **8.00 g of oxygen gas**. How many O₂ molecules is that?

A. 4.15×10^{-25}

B. 0.250

C. 1.51×10^{23}

D. 1.54×10^{26}

✓ Practice 3: answer C

O_2 : $2(16.00) = 32.00 \text{ g/mol}$

given: 8.00 g O_2 · wanted: molecules of O_2

$$8.00 \text{ g } \cancel{\text{O}_2} \times \frac{1 \cancel{\text{ mol } \text{O}_2}}{32.00 \text{ g } \cancel{\text{O}_2}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \cancel{\text{ mol } \text{O}_2}} = 1.51 \times 10^{23} \text{ molecules (answer C)}$$

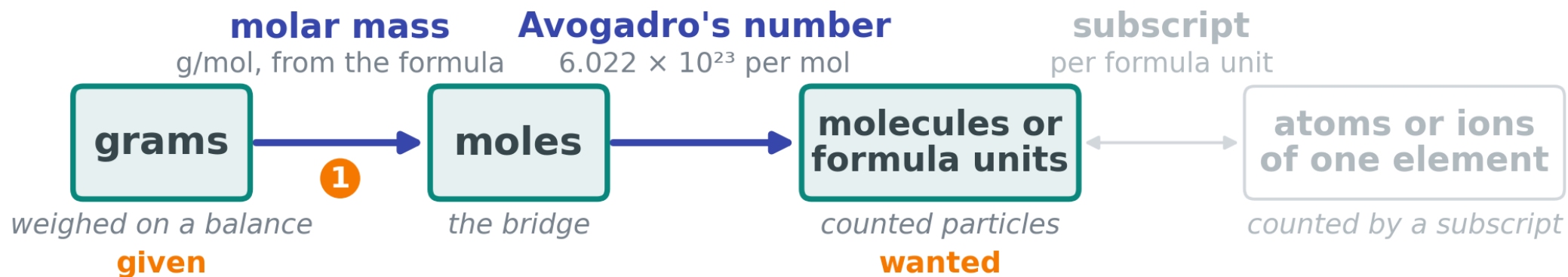
B stopped at moles: $8.00 / 32.00 = 0.250$ counts moles; writing molecules on it does not finish the conversion. D used the factor written 32.00 g over 1 mol : $8.00 \times 32.00 \times 6.022 \times 10^{23} = 1.54 \times 10^{26}$. A used Avogadro's number upside down: $0.250 \div (6.022 \times 10^{23}) = 4.15 \times 10^{-25}$, a fraction of one molecule.

🧠 **Sense check:** 8.00 g is a quarter of a mole of O_2 , and a quarter of 6.022×10^{23} is about 1.5×10^{23} . ✓

✓ Practice 3: the route on the map

O_2 : 32.00 g/mol

given: 8.00 g O_2 · found: 1.51×10^{23} molecules



- ① find the molar mass
- ② write the route: one factor per arrow
- ③ given unit in the denominator
- ④ compute and sense-check

Sense check: Two arrows, two factors. B stopped in the middle box, at moles; A crossed the second arrow with Avogadro's number upside down. ✓



Worked example 4: grams to atoms of an element

Step 1 • Find the molar mass

$$\text{H}_2\text{O}: 2(1.008) + 16.00 = 18.02 \text{ g/mol}$$

given: 25.0 g H₂O • wanted: atoms of H

A glass holds **25.0 g of water**. How many hydrogen **atoms** does it hold?

The question asks for atoms of one element, so the route runs past molecules. The formula's subscript supplies the last equality: 1 H₂O molecule contains 2 H atoms.

✓ Worked example 4: solution

H₂O: 18.02 g/mol · 2 H atoms per molecule

given: 25.0 g H₂O · wanted: atoms of H

Step 2 · Write the route

Step 3 · Pick the orientation that cancels

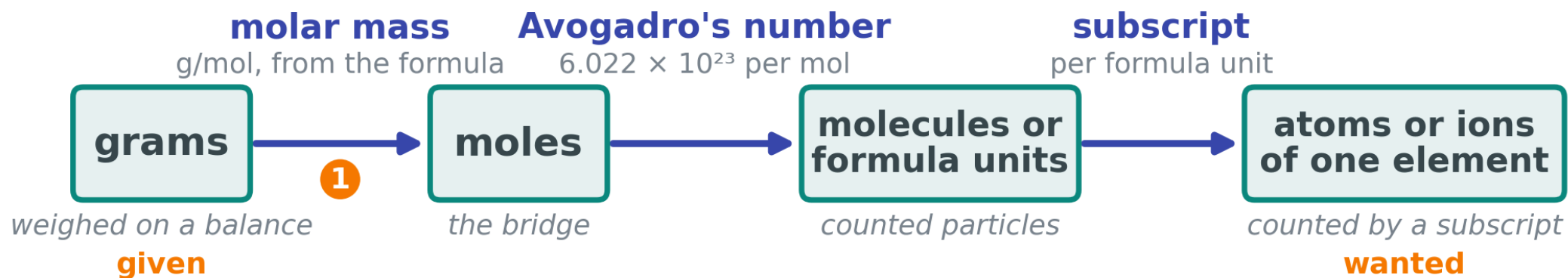
Step 4 · Compute and sense-check

$$25.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol}}{18.02 \text{ g H}_2\text{O}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \times \frac{2 \text{ H atoms}}{1 \text{ molecule}} = 1.67 \times 10^{24} \text{ H atoms}$$

 **Sense check:** The chain passes 8.35×10^{23} molecules on the way. Stopping there and writing atoms undercounts by half; dividing by 2 instead gives 4.18×10^{23} . Two H ride in every molecule, so the atoms outnumber the molecules: 1.67×10^{24} . ✓

✓ Worked example 4: the route on the map

H_2O : $18.02 \text{ g/mol} \cdot 2 \text{ H atoms per molecule}$
 given: $25.0 \text{ g H}_2\text{O}$ · found: $1.67 \times 10^{24} \text{ H atoms}$



- ① find the molar mass
- ② write the route: one factor per arrow
- ③ given unit in the denominator
- ④ compute and sense-check

Sense check: Three arrows: molar mass, Avogadro's number, then the subscript, 2 H atoms per molecule. Atoms of one element sit one box past molecules. ✓



Practice 4

glucose, $\text{C}_6\text{H}_{12}\text{O}_6$

given: 1.20×10^{24} C atoms · wanted: g of glucose

An analysis of a glucose sample counts 1.20×10^{24} carbon atoms. What is the mass of the sample, in grams?

- A. 359
- B. 0.332
- C. 2.15×10^3
- D. 59.8

✓ Practice 4: answer D

$$\text{C}_6\text{H}_{12}\text{O}_6: 6(12.01) + 12(1.008) + 6(16.00) = 180.16 \text{ g/mol}$$

6 C atoms per molecule · given: 1.20×10^{24} C atoms · wanted: g of glucose

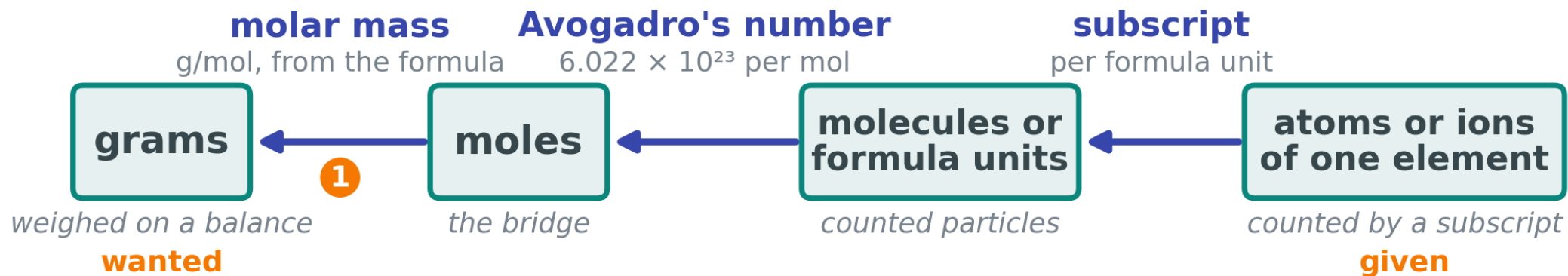
$$1.20 \times 10^{24} \text{ C atoms} \times \frac{1 \text{ molecule}}{6 \text{ C atoms}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \times \frac{180.16 \text{ g}}{1 \text{ mol}} = 59.8 \text{ g (answer D)}$$

A counted every C atom as a molecule: $1.20 \times 10^{24} \div 6.022 \times 10^{23} = 1.99 \text{ mol}$, $\times 180.16 = 359$. B stopped at moles: 0.332 mol of glucose, never converted to grams. C flipped the subscript factor: $\times 6$ instead of $\div 6$ gives 11.96 mol and $2.15 \times 10^3 \text{ g}$.

 **Sense check:** Six C per molecule means a sixth as many molecules as C atoms: 2.00×10^{23} molecules, a third of a mole, a third of 180 g. ✓

✓ Practice 4: the route on the map

$\text{C}_6\text{H}_{12}\text{O}_6$: 180.16 g/mol · 6 C atoms per molecule
 given: 1.20×10^{24} C atoms · found: 59.8 g



- ① find the molar mass
- ② write the route: one factor per arrow
- ③ given unit in the denominator
- ④ compute and sense-check

Sense check: The route runs right to left: the subscript first, then Avogadro's number, then the molar mass. Each factor carries the unit it cancels in its denominator. ✓



Practice 5

$\text{Al}_2(\text{SO}_4)_3$: aluminum sulfate

Al 26.98 · S 32.07 · O 16.00 · sulfate ion: SO_4^{2-}

Water plants add aluminum sulfate to settle fine mud out of drinking water. How many sulfate ions are in a **15.0 g** sample of $\text{Al}_2(\text{SO}_4)_3$?

- A. 5.28×10^{22}
- B. 7.92×10^{22}
- C. 6.02×10^{22}
- D. 2.64×10^{22}
- E. 1.32×10^{23}

✓ Practice 5: answer B

$$\text{Al}_2(\text{SO}_4)_3: 2(26.98) + 3(32.07) + 12(16.00) = 342.17 \text{ g/mol}$$

2 Al · 3 S · 12 O · 3 SO₄²⁻ per formula unit (f.u.) · given: 15.0 g · wanted: SO₄²⁻ ions

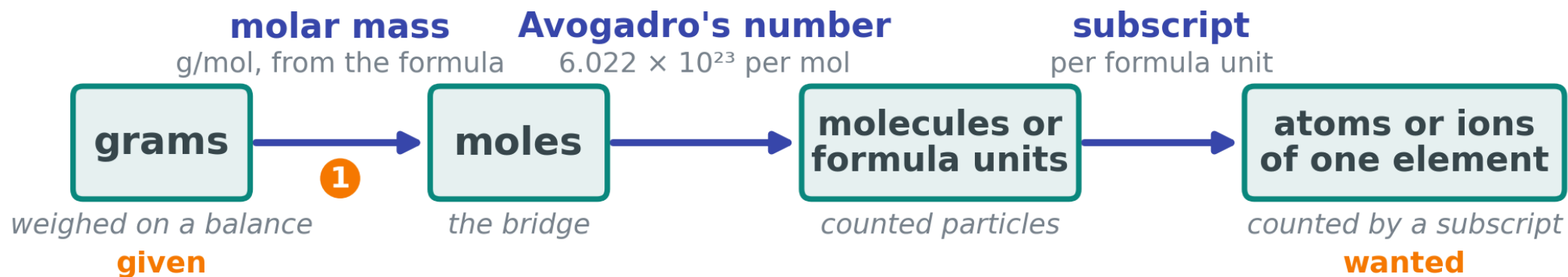
$$15.0 \text{ g} \times \frac{1 \text{ mol}}{342.17 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ f.u.}}{1 \text{ mol}} \times \frac{3 \text{ SO}_4^{2-}}{1 \text{ f.u.}} = 7.92 \times 10^{22} \text{ SO}_4^{2-} \text{ (answer B)}$$

D stopped at formula units: 2.64×10^{22} . A used aluminum's subscript: $\times 2$ gives 5.28×10^{22} . E counted every ion, $2 \text{ Al}^{3+} + 3 \text{ SO}_4^{2-} = 5$ per formula unit: 1.32×10^{23} . C missed the parentheses throughout: Al_2SO_4 at 150.03 g/mol and one sulfate per formula unit: $15.0 \div 150.03 \times 6.022 \times 10^{23} = 6.02 \times 10^{22}$.

🧠 **Sense check:** The sample is 0.0438 mol, or 2.64×10^{22} formula units, and each carries 3 sulfate ions. ✓

✓ Practice 5: the route on the map

$\text{Al}_2(\text{SO}_4)_3$: 342.17 g/mol · 3 SO_4^{2-} per formula unit
 given: 15.0 g · found: 7.92×10^{22} SO_4^{2-}



- ① find the molar mass
- ② write the route: one factor per arrow
- ③ given unit in the denominator
- ④ compute and sense-check

Sense check: Three arrows: molar mass, Avogadro's number, then the subscript on (SO_4). The 3 outside the parentheses enters twice: in the molar mass and in the last factor. ✓



Extra practice: calcium phosphate

$\text{Ca}_3(\text{PO}_4)_2$: calcium phosphate

Ca 40.08 · P 30.97 · O 16.00

A bone-graft paste contains **0.0150 mol** of $\text{Ca}_3(\text{PO}_4)_2$. How many grams of $\text{Ca}_3(\text{PO}_4)_2$ is that?

- A. 4.65
- B. 3.23
- C. 4.84×10^{-5}
- D. 3.69

✓ Extra practice: answer A

$$\text{Ca}_3(\text{PO}_4)_2: 3(40.08) + 2(30.97) + 8(16.00) = 310.18 \text{ g/mol}$$

3 Ca · 2 P · 8 O · given: 0.0150 mol · wanted: g

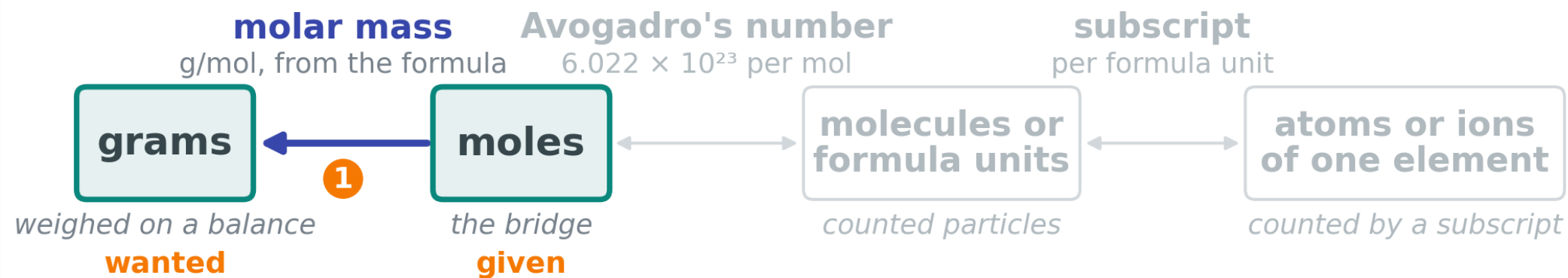
$$0.0150 \text{ mol} \times \frac{310.18 \text{ g}}{1 \text{ mol}} = 4.65 \text{ g (answer A)}$$

B read PO_4 once: $3(40.08) + 30.97 + 4(16.00) = 215.21 \text{ g/mol}$, and $0.0150 \times 215.21 = 3.23$. D applied the 2 to P but not to the O inside: 246.18 g/mol gives 3.69. C divided by the molar mass: $0.0150 \div 310.18 = 4.84 \times 10^{-5}$.

 **Sense check:** The 2 outside the parentheses doubles P and O alike: 2 P, 8 O. A small fraction of a mole of a 310 g/mol solid weighs a few grams. ✓

✓ Extra practice: the route on the map

$\text{Ca}_3(\text{PO}_4)_2$: 310.18 g/mol
given: 0.0150 mol · found: 4.65 g



- ① find the molar mass
- ② write the route: one factor per arrow
- ③ given unit in the denominator
- ④ compute and sense-check

🧠 **Sense check:** One arrow, moles to grams. The parentheses act in Step 1: 2 P and 8 O enter the molar mass. ✓

Check yourself

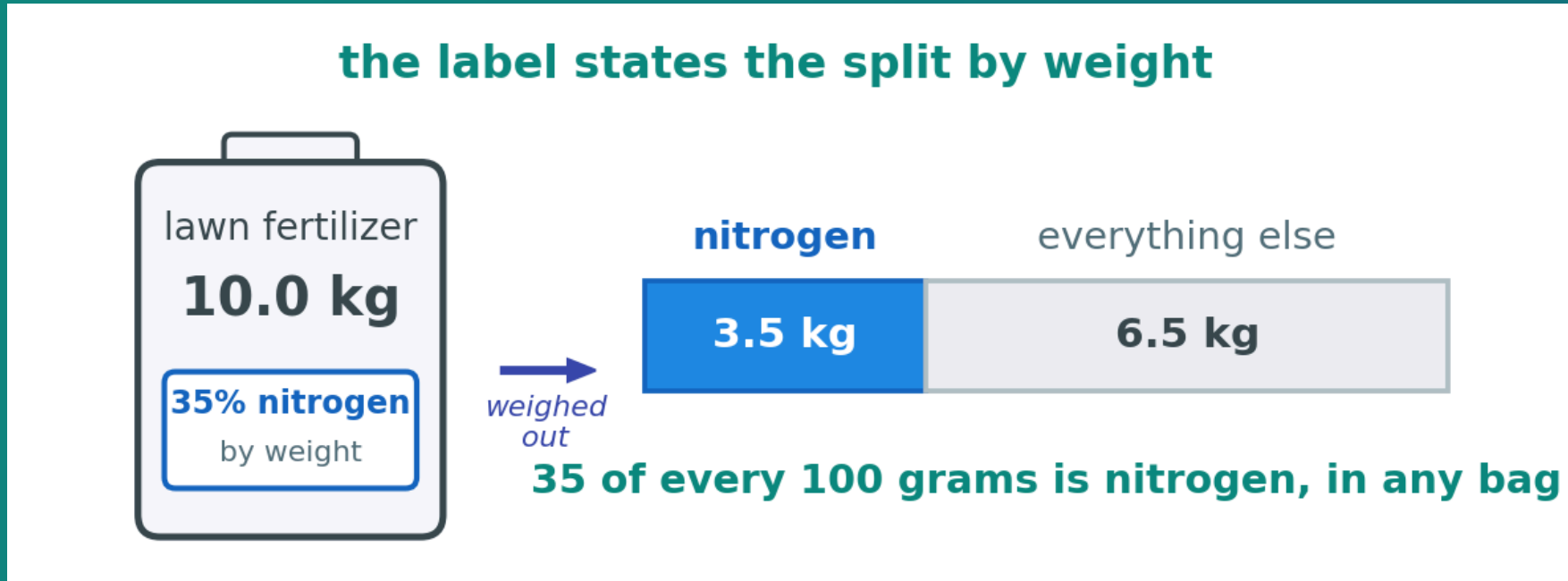
1. From **1 mol Cu = 63.55 g Cu**, write both conversion factors. Which one converts 12.7 g of copper to moles, and which converts 0.200 mol of copper to grams?
2. One mole of CO₂ weighs 44.01 g. Is a 22 g sample more or less than one mole? More or fewer than 6.022×10^{23} molecules?

Molar mass also splits a compound into its elements' shares: 12.01 g of every 44.01 g of CO₂ is carbon. That share, written as a percent, is the **percent composition**.

2 · Percent Composition

Calculate the mass percent of any element in a compound from its formula, and confirm the result with the percents-total-100 check.

The number on the bag



Fertilizer bags state how much of the weight is nitrogen: 35%, so a 10.0-kg bag carries 3.5 kg of nitrogen. Every bag, any size, keeps the split.

Percent from measured masses

3.67 g chalcopyrite: 1.27 g Cu · 1.12 g Fe · 1.28 g S

part: one element's mass · whole: the 3.67 g sample · $1.27 + 1.12 + 1.28 = 3.67$ g

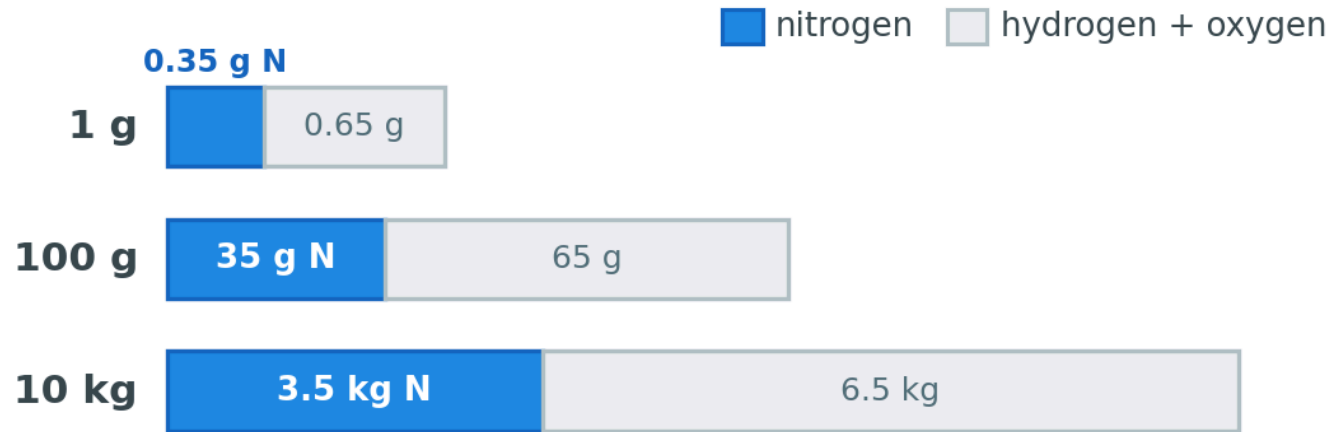
A lab weighed each element in a mineral sample. Each element's share is part over whole, $\times 100$.

$$\% \text{Cu} = \frac{1.27 \text{ g Cu}}{3.67 \text{ g sample}} \times 100 = \mathbf{34.6\% \text{ Cu}}$$

Iron 30.5%, sulfur 34.9%: the three total 100.0%. Every 100 g of chalcopyrite holds 34.6 g of copper.

The formula fixes the mass split

ammonium nitrate, NH_4NO_3 : three sample sizes



the nitrogen share never changes: 35.00% of the mass in every sample

A compound's formula fixes its recipe by mass. Ammonium nitrate is NH_4NO_3 in every crystal, so nitrogen's share of the mass is the same in every sample, of any size.

Mass percent of an element

$$\text{mass \% of an element} = \frac{\text{element mass in one mole}}{\text{molar mass}} \times 100$$

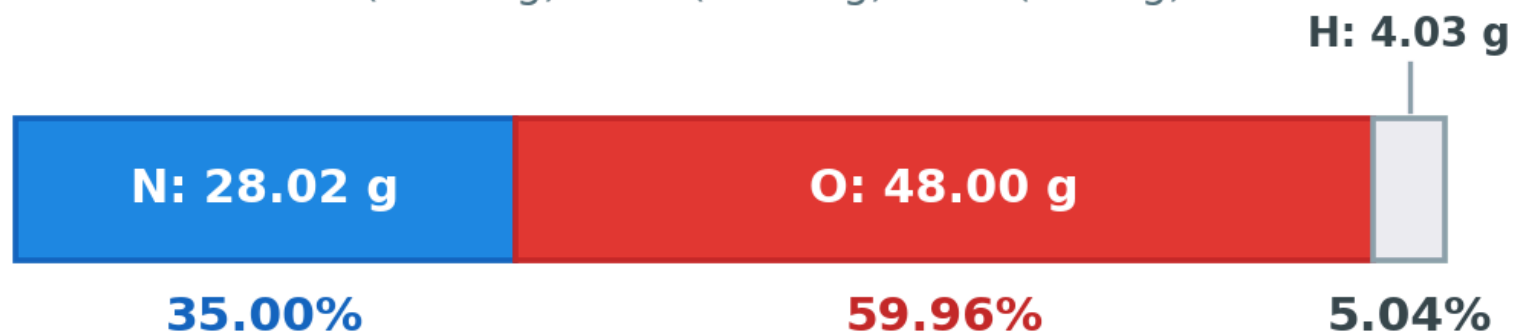
$$\text{element mass in one mole} = \text{atomic mass} \times \text{subscript}$$

One mole of the compound is the sample: the molar mass is the whole, and the element's atoms supply the part. The part over the whole, $\times 100$, is the percent.

The percents total 100

one mole of NH_4NO_3 : 80.05 g

2 N (28.02 g) · 3 O (48.00 g) · 4 H (4.03 g)



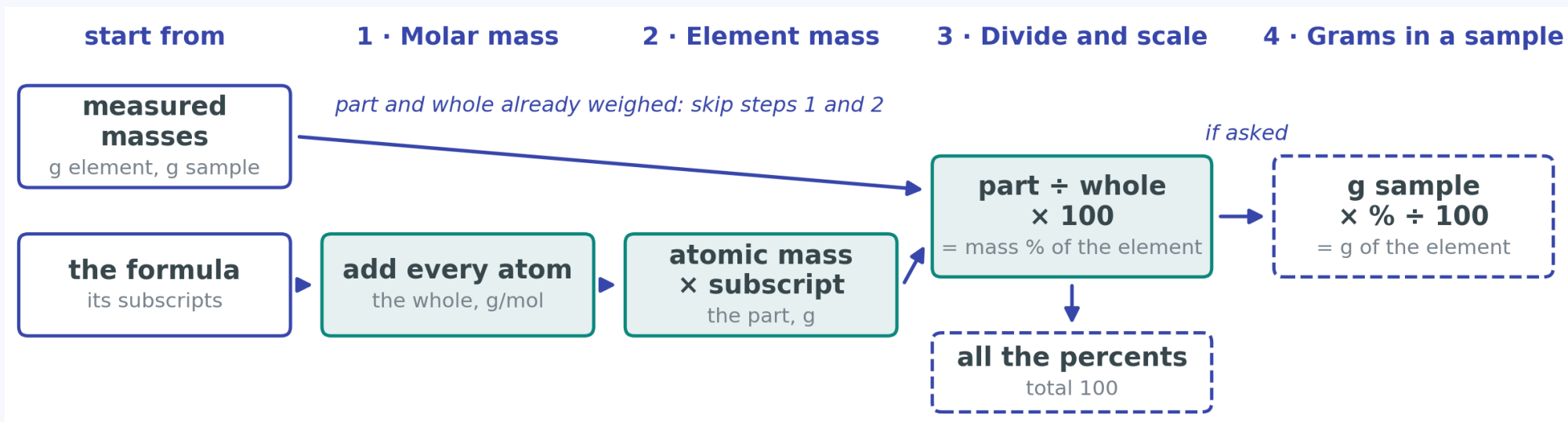
35.00 + 59.96 + 5.04 = 100.00: the shares cover the whole

Each element takes its share of the molar mass, and the shares cover the whole. The percents of all elements total 100: the built-in check on the arithmetic.



The method

1. **Molar mass:** add the mass of every atom in the formula.
2. **Element mass:** atomic mass $\times$ subscript, for the element asked about.
3. **Divide and scale:** element mass over molar mass, $\times 100$.
4. **Grams in a sample:** apply the percent.





Guided example: calcium in limestone

CaCO_3 : calcium carbonate

Ca 40.08 · C 12.01 · O 16.00 g/mol · wanted: mass % Ca

Limestone and chalk are calcium carbonate. Find the mass percent of calcium in CaCO_3 .

The formula supplies both numbers. The molar mass is the whole; calcium's atoms are the part.

✓ Guided example: solution

CaCO₃: calcium carbonate

Ca 40.08 · C 12.01 · O 16.00 g/mol · wanted: mass % Ca

Step 1 · Molar mass

CaCO₃: 40.08 + 12.01 + 3(16.00) = 100.09 g/mol

1 Ca · 1 C · 3 O · the whole

Step 2 · Element mass

Step 3 · Divide and scale

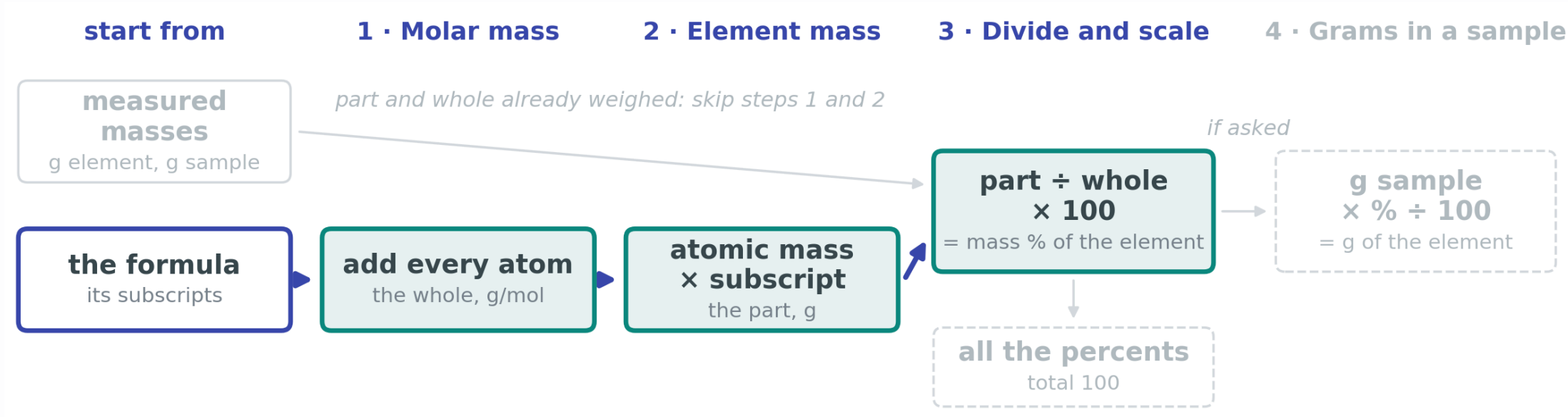
$$\% \text{Ca} = \frac{40.08 \text{ g Ca}}{100.09 \text{ g CaCO}_3} \times 100 = \mathbf{40.04\% \text{ Ca}}$$

 **Sense check:** The molar mass is close to 100, so the percent lands close to the 40.08 g of calcium itself. ✓

✓ Guided example: the route on the map

CaCO₃: calcium carbonate

100.09 g/mol · 40.08 g Ca per mole · found: 40.04% Ca



 **Sense check:** The formula supplies the part and the whole: steps 1, 2 and 3 in order. No sample mass was given, so step 4 stays unlit. ✓

 Practice 1

galena, a lead ore: 12.99 g Pb · 2.01 g S
given: 15.00 g sample · 12.99 g Pb · 2.01 g S · wanted: mass % S

Analysis of a **15.00 g** sample of galena finds **12.99 g** of lead and **2.01 g** of sulfur. What is the mass percent of sulfur?

- A. 15.5
- B. 13.4
- C. 86.6
- D. 746

✓ Practice 1: answer B

galena, a lead ore: 12.99 g Pb · 2.01 g S

given: 15.00 g sample · wanted: mass % S · part: 2.01 g S · whole: 15.00 g sample

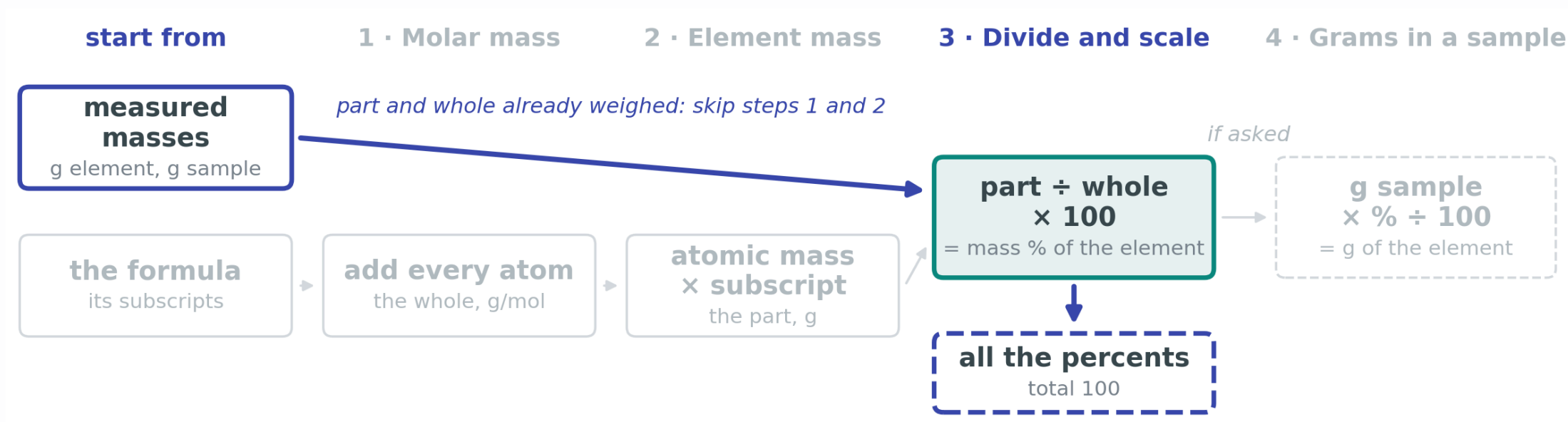
$$\%S = \frac{2.01 \text{ g S}}{15.00 \text{ g sample}} \times 100 = \mathbf{13.4\% \text{ S (answer B)}}$$

A divided by the lead instead of the whole sample: $2.01 \div 12.99 \times 100 = 15.5$. C is lead's share: $12.99 \div 15.00 \times 100 = 86.6$. D flipped part and whole: $15.00 \div 2.01 \times 100 = 746$, a share above 100.

🧠 **Sense check:** $13.4 + 86.6 = 100.0$: sulfur and lead cover the whole sample. ✓

✓ Practice 1: the route on the map

galena, a lead ore: 12.99 g Pb · 2.01 g S
15.00 g sample · found: 13.4% S



Sense check: The lab weighed the part and the whole, so no molar mass is needed. One division, then the 100% check. ✓



Worked example 1: potassium chloride

KCl: potassium chloride

K 39.10 · Cl 35.45 g/mol · wanted: mass % of each element

Salt substitute for low-sodium diets is potassium chloride. Find the mass percent of potassium and of chlorine.

✓ Worked example 1: solution

Step 1 • Molar mass

$$\text{KCl: } 39.10 + 35.45 = 74.55 \text{ g/mol}$$

one K + one Cl · every atom in the formula · wanted: mass % of each element

Step 2 • Element mass

Step 3 • Divide and scale

$$\%K = \frac{39.10 \text{ g K}}{74.55 \text{ g KCl}} \times 100 = \mathbf{52.45\% K}$$

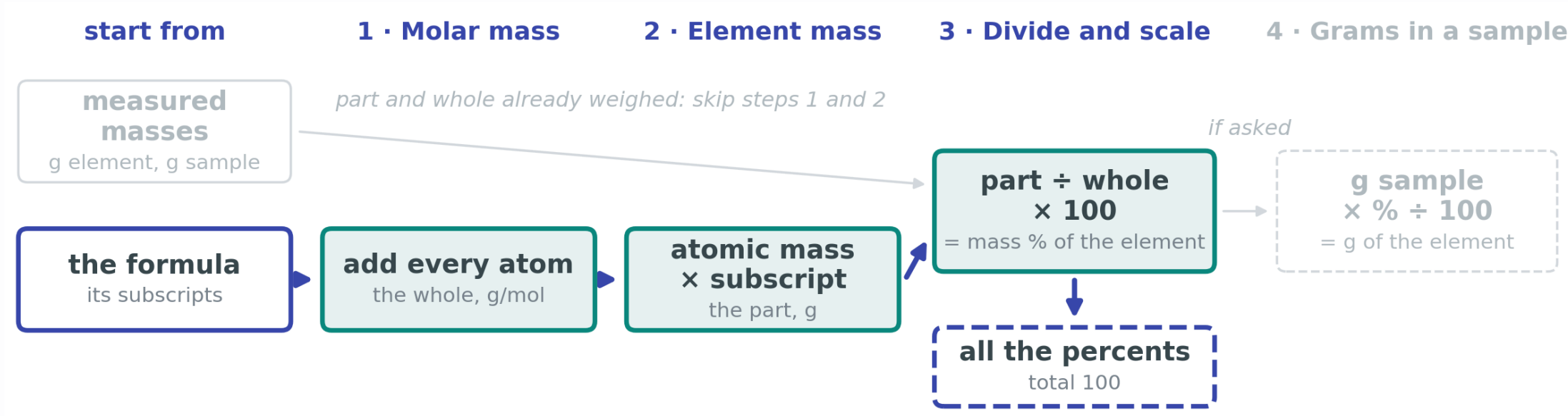
$$\%Cl = \frac{35.45 \text{ g Cl}}{74.55 \text{ g KCl}} \times 100 = \mathbf{47.55\% Cl}$$

🧠 **Sense check:** $52.45 + 47.55 = 100.00$: the two shares cover the whole compound. ✓

✓ Worked example 1: the route on the map

$$\text{KCl: } 39.10 + 35.45 = 74.55 \text{ g/mol}$$

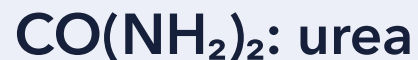
found: 52.45% K · 47.55% Cl



 **Sense check:** The formula supplies the part and the whole. With every element found, the 100% check closes the route. ✓



Worked example 2: urea



C 12.01 · O 16.00 · N 14.01 · H 1.008 g/mol · wanted: mass % N

Urea is a solid nitrogen fertilizer. Find the mass percent of nitrogen.

A common first attempt: nitrogen's atomic mass is 14.01, so take 14.01 g of N per mole.
Test it against the formula.

✓ Worked example 2: solution

Step 1 • Molar mass

urea $\text{CO}(\text{NH}_2)_2$: $12.01 + 16.00 + 2(14.01) + 4(1.008) = 60.06 \text{ g/mol}$
1 C · 1 O · 2 N · 4 H · the outside 2 multiplies the parentheses · wanted: mass % N

Step 2 • Element mass = atomic mass × subscript

$$2 \times 14.01 = 28.02 \text{ g N in one mole}$$

Step 3 • Divide and scale

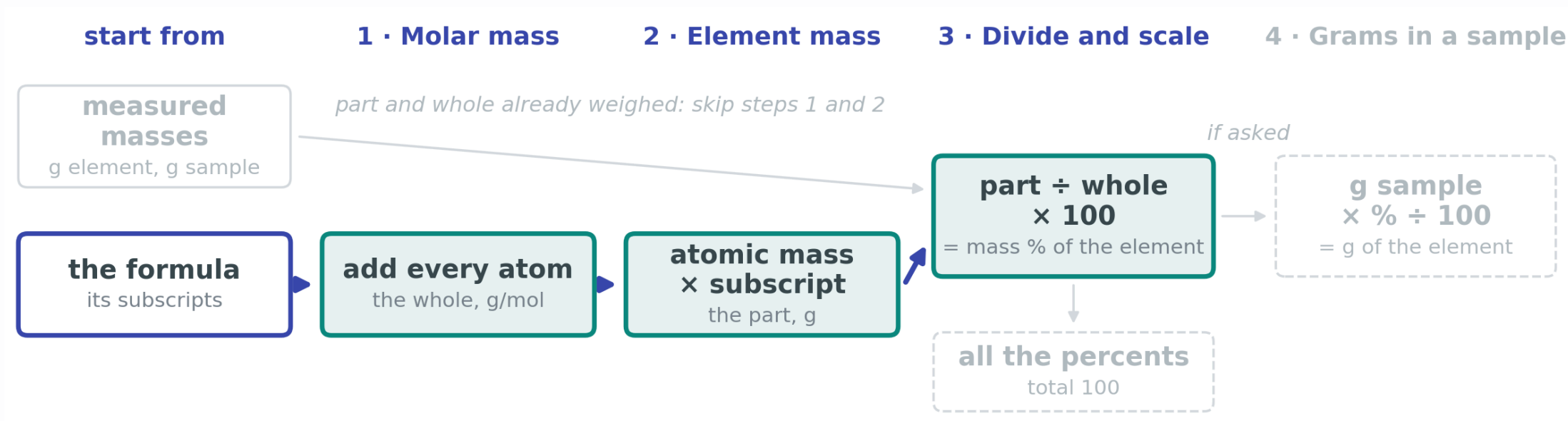
$$\% \text{N} = \frac{28.02 \text{ g N}}{60.06 \text{ g CO}(\text{NH}_2)_2} \times 100 = 46.65\% \text{ N}$$

🧠 **Sense check:** Urea is labeled 46% nitrogen; 46.65 matches. One N is half: $14.01 \div 60.06 \times 100 = 23.33\% \times$

✓ Worked example 2: the route on the map

urea $\text{CO}(\text{NH}_2)_2$: 60.06 g/mol

$2 \times 14.01 = 28.02$ g N per mole · found: 46.65% N



🧠 **Sense check:** Step 2 carries the outside 2: two N atoms, 28.02 g, go on top. ✓

Take-home: the subscript multiplies the mass

one N: $14.01 \div 60.06 \times 100 = 23.33\%$ ✗
CO(NH2)2 · the outside 2 doubles the NH2 group: 2 N, 4 H

two N: $28.02 \div 60.06 \times 100 = 46.65\%$ ✓
element mass = atomic mass × subscript

The numerator carries every atom of the element in the formula. Read every subscript before the division.

Your turn: sulfur dioxide

$$\text{SO}_2: 32.07 + 2(16.00) = 64.07 \text{ g/mol}$$

$$\text{S } 32.07 \cdot \text{O } 16.00 \text{ g/mol} \cdot \text{wanted: mass \% O}$$

Sulfur dioxide forms when coal burns. Fill the element mass from the subscript, then complete the percent:

$$\% \text{O} = \frac{\text{ } \text{g O}}{64.07 \text{ g SO}_2} \times 100 = \text{ } \% \text{O}$$

$$\% \text{O} = \frac{32.00 \text{ g O}}{64.07 \text{ g SO}_2} \times 100 = \mathbf{49.95\% \text{ O}}$$

 **Sense check:** Sulfur takes the other half: $32.07 \div 64.07 \times 100 = 50.05\% \text{ S}$, and $49.95 + 50.05 = 100.00$. ✓

⚠ Where this goes wrong



$$\% \text{O in SO}_2 = 49.95 \cdot \% \text{N in NH}_4\text{NO}_3 = 35.00$$

Counting atoms instead of mass. Two of SO₂'s three atoms are oxygen: $2 \div 3 \times 100 = 66.67\%$ of the atoms. Mass percent weighs them: one S outweighs one O, 32.07 vs 16.00, and the mass share is 49.95%. Atoms are counted; percent composition is weighed.

Reporting everything except the element. Asked for %N in NH₄NO₃: $52.03 \div 80.05 \times 100 = 65.00\%$ is the share of the H and O. The asked-for element's mass goes on top: $28.02 \div 80.05 \times 100 = 35.00\%$.

Dividing by another element's mass. 32.00 g O over 32.07 g S gives 99.78%. The whole compound belongs underneath: $32.00 \div 64.07 \times 100 = 49.95\%$. A share near 100 from half the mass fails the sum check.



Practice 2

Mg₃N₂: magnesium nitride

Mg 24.31 · N 14.01 g/mol · wanted: mass % N

Magnesium nitride is a yellow-green ionic solid. What is the mass percent of nitrogen in Mg₃N₂?

- A. 72.24
- B. 40.0
- C. 38.42
- D. 27.76

✓ Practice 2: answer D

$$\text{Mg}_3\text{N}_2: 3(24.31) + 2(14.01) = 100.95 \text{ g/mol}$$

$$3 \text{ Mg} \cdot 2 \text{ N} \cdot \text{element mass: } 2 \times 14.01 = 28.02 \text{ g N per mole}$$

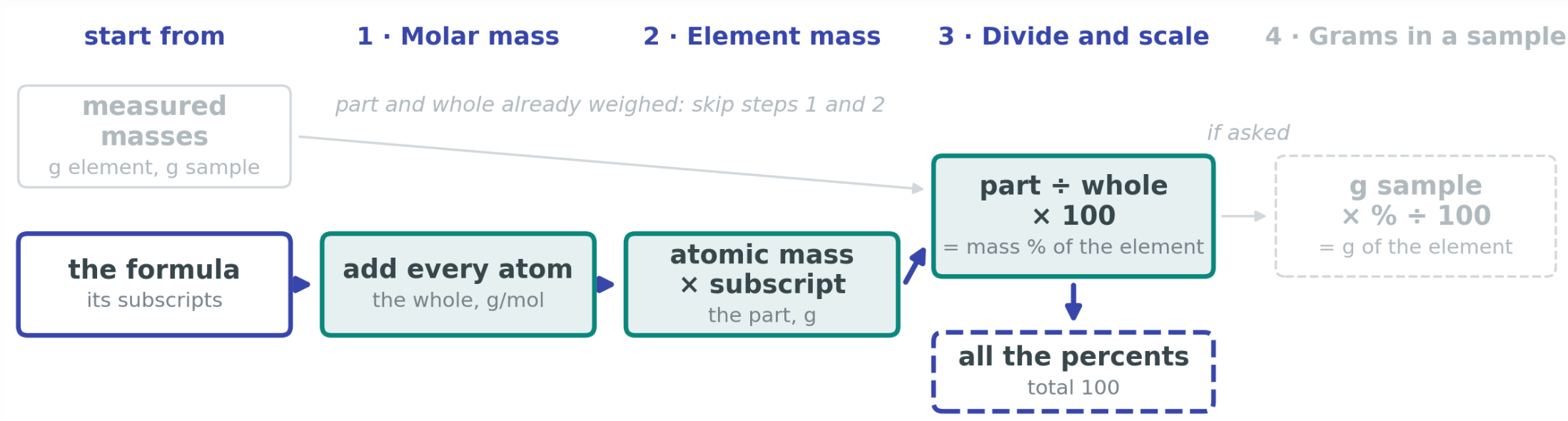
$$\% \text{N} = \frac{28.02 \text{ g N}}{100.95 \text{ g Mg}_3\text{N}_2} \times 100 = \mathbf{27.76\% \text{ N (answer D)}}$$

A is magnesium's share: $72.93 \div 100.95 \times 100 = 72.24$. B counted atoms: 2 of 5 atoms is 40.0%, but an N atom weighs less than an Mg atom. C divided by magnesium's mass instead of the whole: $28.02 \div 72.93 \times 100 = 38.42$.

🧠 **Sense check:** $27.76 + 72.24 = 100.00$: nitrogen and magnesium cover the whole compound. ✓

✓ Practice 2: the route on the map

Mg_3N_2 : 100.95 g/mol
28.02 g N per mole · found: 27.76% N



Sense check: The formula supplies the part and the whole, and magnesium's 72.24% closes the 100% check. ✓



Worked example 3: grams of an element in a sample

CaCO₃: calcium carbonate

Ca 40.08 · C 12.01 · O 16.00 g/mol · given: 1.5×10^4 g CaCO₃ · wanted: g Ca

A quarry block of pure limestone, CaCO₃, has a mass of 1.5×10^4 g. How many grams of calcium does it hold?

A percent is grams of element per 100 g of compound. As a conversion factor, it carries the sample to the element.

✓ Worked example 3: solution

CaCO_3 : calcium carbonate

given: $1.5 \times 10^4 \text{ g CaCO}_3$ · wanted: g Ca

Step 1 · Molar mass

Step 2 · Element mass

Step 3 · Divide and scale

$$\% \text{Ca} = 40.08 \div 100.09 \times 100 = 40.04\% \text{ Ca}$$

CaCO_3 : 100.09 g/mol · 1 Ca: 40.08 g per mole · 40.04 g Ca in every 100 g CaCO_3

Step 4 · Grams in a sample

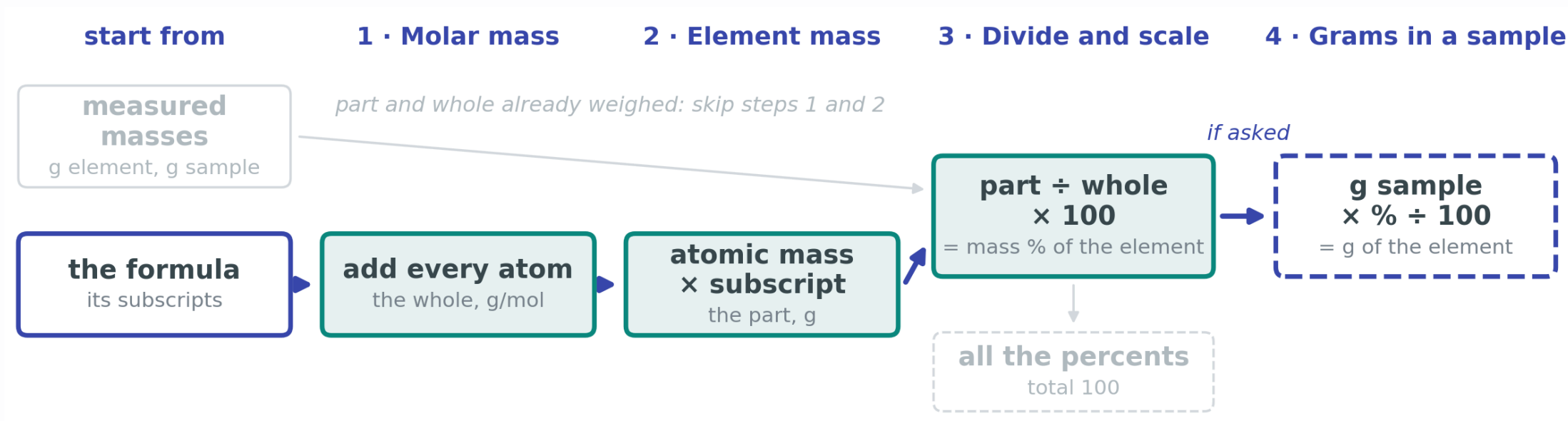
$$1.5 \times 10^4 \text{ g } \cancel{\text{CaCO}_3} \times \frac{40.04 \text{ g Ca}}{100 \text{ g } \cancel{\text{CaCO}_3}} = 6.0 \times 10^3 \text{ g Ca}$$

 **Sense check:** Calcium is about 40% of the block: $0.40 \times 1.5 \times 10^4 \text{ g} = 6.0 \times 10^3 \text{ g}$ ✓. The flipped factor, $1.5 \times 10^4 \times 100 \div 40.04 = 3.7 \times 10^4 \text{ g}$, claims more calcium than limestone ✗

✓ Worked example 3: the route on the map

CaCO_3 : 40.04% Ca

given: $1.5 \times 10^4 \text{ g CaCO}_3$ · found: $6.0 \times 10^3 \text{ g Ca}$



🧠 **Sense check:** Steps 1 to 3 give the percent. Step 4 uses it as a factor on the sample mass. ✓



Practice 3

K_2CO_3 : potassium carbonate

K 39.10 · C 12.01 · O 16.00 g/mol · given: 250. g bag

Glassmakers melt potassium carbonate with sand. How many grams of potassium are in a **250. g** bag of K_2CO_3 ?

- A. 141
- B. 70.7
- C. 109
- D. 56.6
- E. 83.3

✓ Practice 3: answer A

$$\text{K}_2\text{CO}_3: 2(39.10) + 12.01 + 3(16.00) = 138.21 \text{ g/mol}$$

$$2 \text{ K}: 78.20 \text{ g K per mole} \cdot \%K = 78.20 \div 138.21 \times 100 = 56.58\%$$

$$250. \text{ g } \cancel{\text{K}_2\text{CO}_3} \times \frac{56.58 \text{ g K}}{100 \cancel{\text{ g } \text{K}_2\text{CO}_3}} = 141 \text{ g K (answer A)}$$

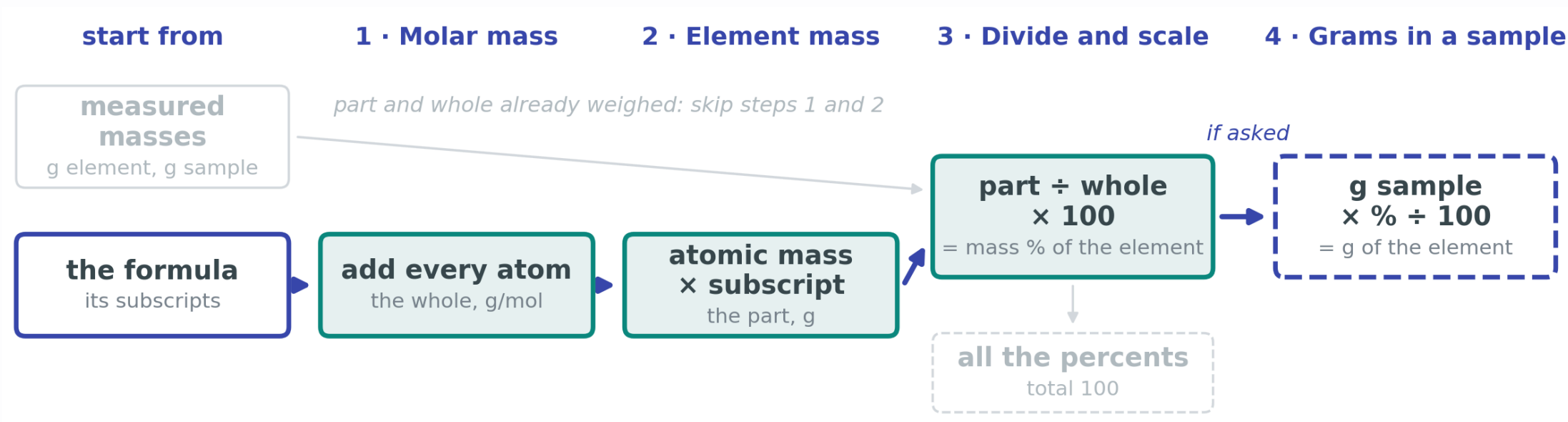
B used one K: $250. \times 39.10 \div 138.21 = 70.7$. C is the carbon and oxygen: $250. \times 60.01 \div 138.21 = 109$. D stopped at the percent, 56.6, before using it on the bag. E counted atoms: 2 of 6, $250. \times 2 \div 6 = 83.3$.

 **Sense check:** Potassium is a bit over half the compound, and 141 g is a bit over half of 250. g. ✓

✓ Practice 3: the route on the map

K_2CO_3 : 56.58% K

given: 250. g bag · found: 141 g K



🧠 **Sense check:** Steps 1 to 3 give the percent, and step 4 applies it to the bag. ✓



Practice 4

$(\text{NH}_4)_3\text{PO}_4$: ammonium phosphate

given: 5.00 kg sack · 40.0% $(\text{NH}_4)_3\text{PO}_4$ by mass · wanted: g of N

A **5.00 kg** sack of lawn feed is **40.0%** ammonium phosphate by mass; the rest is filler with no nitrogen. How many grams of nitrogen does the sack deliver?

- A. 188
- B. 2.00×10^3
- C. 564
- D. 7.09×10^3
- E. 1.41×10^3

✓ Practice 4: answer C

$$(\text{NH}_4)_3\text{PO}_4: 3(14.01) + 12(1.008) + 30.97 + 4(16.00) = 149.10 \text{ g/mol}$$

$$\text{the outside 3 triples NH}_4 \cdot 3 \times 14.01 = 42.03 \text{ g N per mole} \cdot \%N = 42.03 \div 149.10 \times 100 = 28.19\%$$

$$5.00 \times 10^3 \text{ g feed} \times \frac{40.0 \text{ g } (\text{NH}_4)_3\text{PO}_4}{100 \text{ g feed}} \times \frac{42.03 \text{ g N}}{149.10 \text{ g } (\text{NH}_4)_3\text{PO}_4} = 564 \text{ g N (answer C)}$$

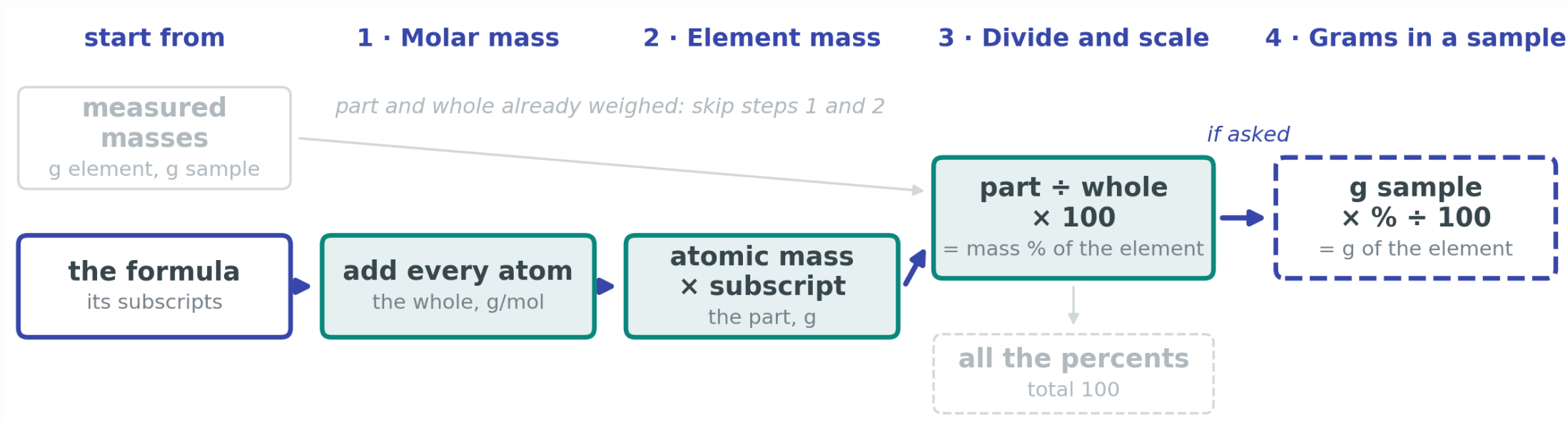
A counted one N of three: $2.00 \times 10^3 \times 14.01 \div 149.10 = 188$. B stopped halfway: $2.00 \times 10^3 \text{ g}$ is the ammonium phosphate in the sack, not its nitrogen. D flipped part and whole: $2.00 \times 10^3 \times 149.10 \div 42.03 = 7.09 \times 10^3$, more nitrogen than compound. E skipped the 40.0%: $5.00 \times 10^3 \times 42.03 \div 149.10 = 1.41 \times 10^3$ counts the filler as fertilizer.

 **Sense check:** The sack holds 2.00 kg of the compound, and 28.19% of that is a bit over a quarter: 564 g. ✓

✓ Practice 4: the route on the map



given: 5.00 kg sack · 40.0% $(\text{NH}_4)_3\text{PO}_4$ · found: 564 g N



Sense check: Step 4 runs twice: 40.0% takes the sack to the compound, then 28.19% takes the compound to its nitrogen. ✓

 Practice 5

$\text{Al}_2(\text{SO}_4)_3$: aluminum sulfate

Al 26.98 · S 32.07 · O 16.00 g/mol · wanted: mass % O

Water-treatment plants add aluminum sulfate to settle fine particles out of the water. What is the mass percent of oxygen in $\text{Al}_2(\text{SO}_4)_3$?

- A. 18.70
- B. 42.66
- C. 70.6
- D. 43.89
- E. 56.11

✓ Practice 5: answer E

$$\text{Al}_2(\text{SO}_4)_3: 2(26.98) + 3(32.07) + 12(16.00) = 342.17 \text{ g/mol}$$

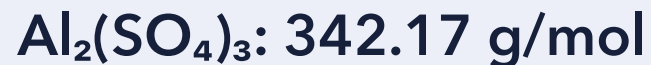
2 Al · 3 S · 12 O · the outside 3 multiplies all 4 O in SO_4 : $3 \times 4 = 12$

$$\% \text{O} = \frac{192.00 \text{ g O}}{342.17 \text{ g Al}_2(\text{SO}_4)_3} \times 100 = \mathbf{56.11\% \text{ O (answer E)}}$$

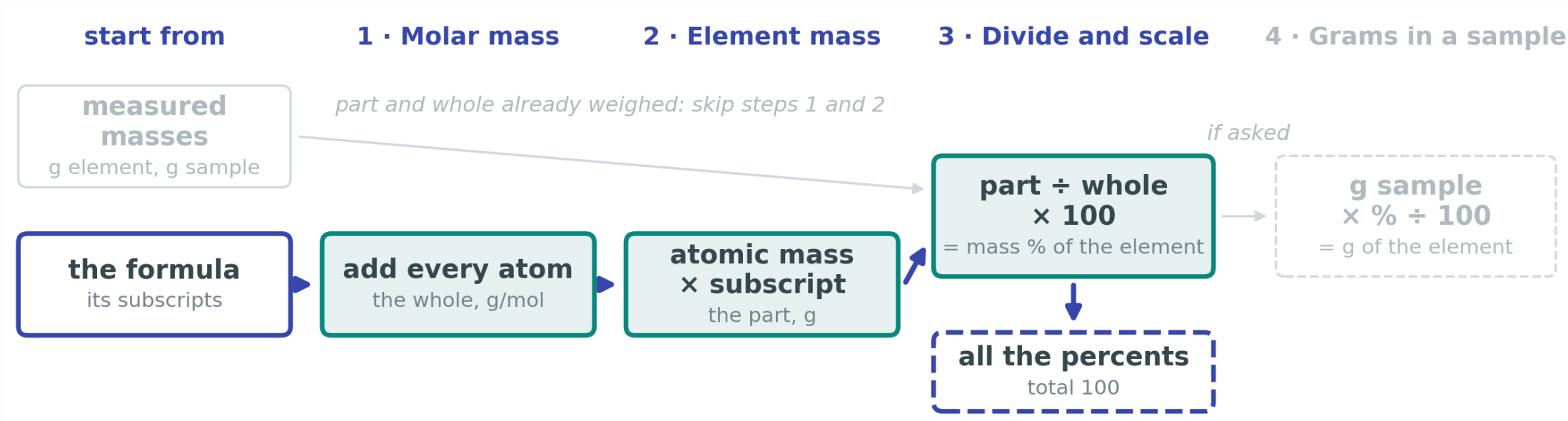
A put 4 O on top but kept the right whole: $64.00 \div 342.17 \times 100 = 18.70$. B read the formula as Al_2SO_4 everywhere: $64.00 \div 150.03 \times 100 = 42.66$. C counted atoms: 12 of 17 = 70.6. D is the aluminum and sulfur: $150.17 \div 342.17 \times 100 = 43.89$.

🧠 **Sense check:** $56.11 + 43.89 = 100.00$: oxygen and the rest cover the whole compound. ✓

✓ Practice 5: the route on the map



$$12 \times 16.00 = 192.00 \text{ g O per mole} \cdot \text{found: } 56.11\% \text{ O}$$



🧠 **Sense check:** Steps 1 and 2 both carry the outside 3: twelve O in the whole and twelve on top. ✓

Check yourself

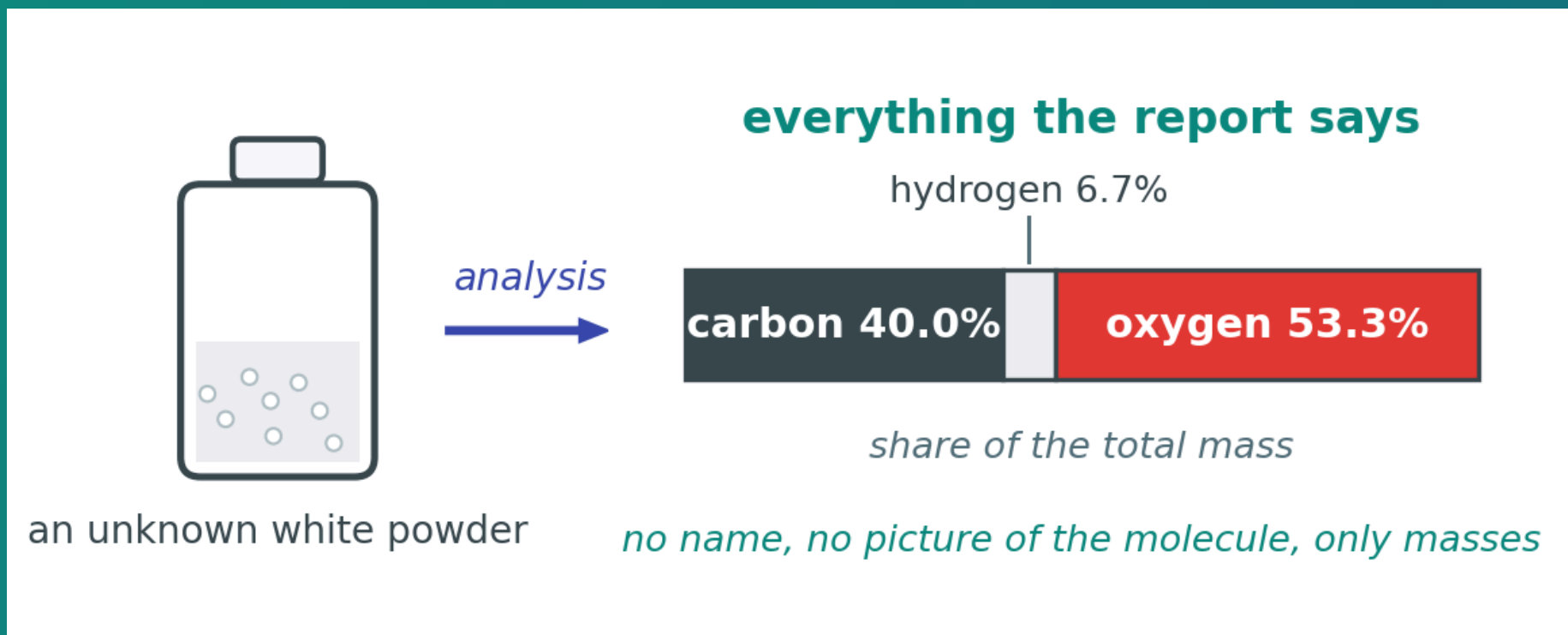
1. A 5.0-g pinch and a 5.0-kg drum hold the same pure compound. Compare their mass percents of each element. What fixes those numbers?
2. Set up %C in glucose, $\text{C}_6\text{H}_{12}\text{O}_6$: which number multiplies 12.01 in the numerator, and what goes in the denominator?

A compound's percents are fixed by its formula. A solution's concentration is set by whoever prepares it, and the working measure counts moles of solute in each liter of solution: the **molarity**.

3 · Empirical & Molecular Formulas

Turn percent-composition data into the empirical formula, then scale it with the molar mass to reach the molecular formula.

An unknown powder, three numbers



A lab receives an unknown white powder. The analysis reports only its makeup by mass: 40.0% carbon, 6.7% hydrogen, 53.3% oxygen. Those three masses pin down its empirical formula.

A formula is a mole ratio

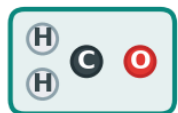


atoms = 6 C : 12 H : 6 O · moles = 6 : 12 : 6, the same ratio · grams = 72.1 : 12.1 : 96.0, not the same

Subscripts count atoms. Moles count atoms in bulk, so a sample's mole ratio equals its atom ratio. Grams do not: a carbon atom has about twelve times the mass of a hydrogen atom.

Molecular formula vs empirical formula

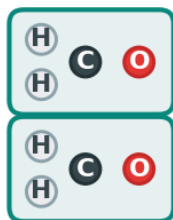
one repeating unit, 1 C : 2 H : 1 O



x 1



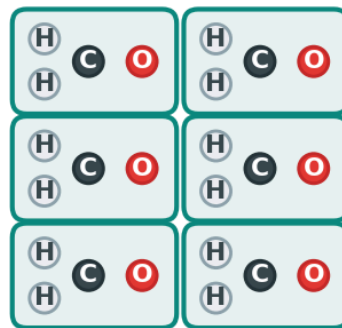
formaldehyde



x 2



acetic acid



x 6



glucose

same smallest ratio, different molecule sizes

The molecular formula counts the atoms in one molecule. The empirical formula reduces that count to the smallest whole-number ratio. Water's H₂O is already smallest. For many compounds the two match.

From molecular to empirical formula



benzene $\div 6$ · acetylene $\div 2$ · hydrogen peroxide $\div 2$ · water: 2 and 1 share no factor, so it stays

Divide every subscript by the largest number that divides them all. When no such number exists, the molecular and empirical formulas match.

Reduce butane, C_4H_{10} .



4 and 10 share a factor of 2 · $\div 4$ would leave $\text{H}_{2.5}$, not a whole number

The 100 g sample

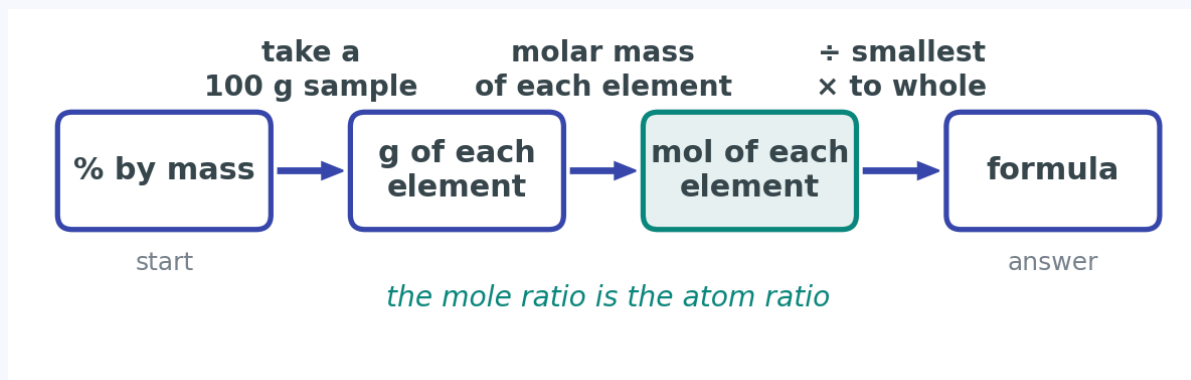
40.0% C $\rightarrow$ 40.0 g C in every 100 g of compound

per 100 g: 40.0 g C $\cdot$ 6.7 g H $\cdot$ 53.3 g O $\cdot$ total 100.0 g $\checkmark$

Percent means parts per hundred. Choosing a 100 g sample turns each percent into the same number of grams. Any sample size gives the same formula; 100 g is the convenient choice.



The method

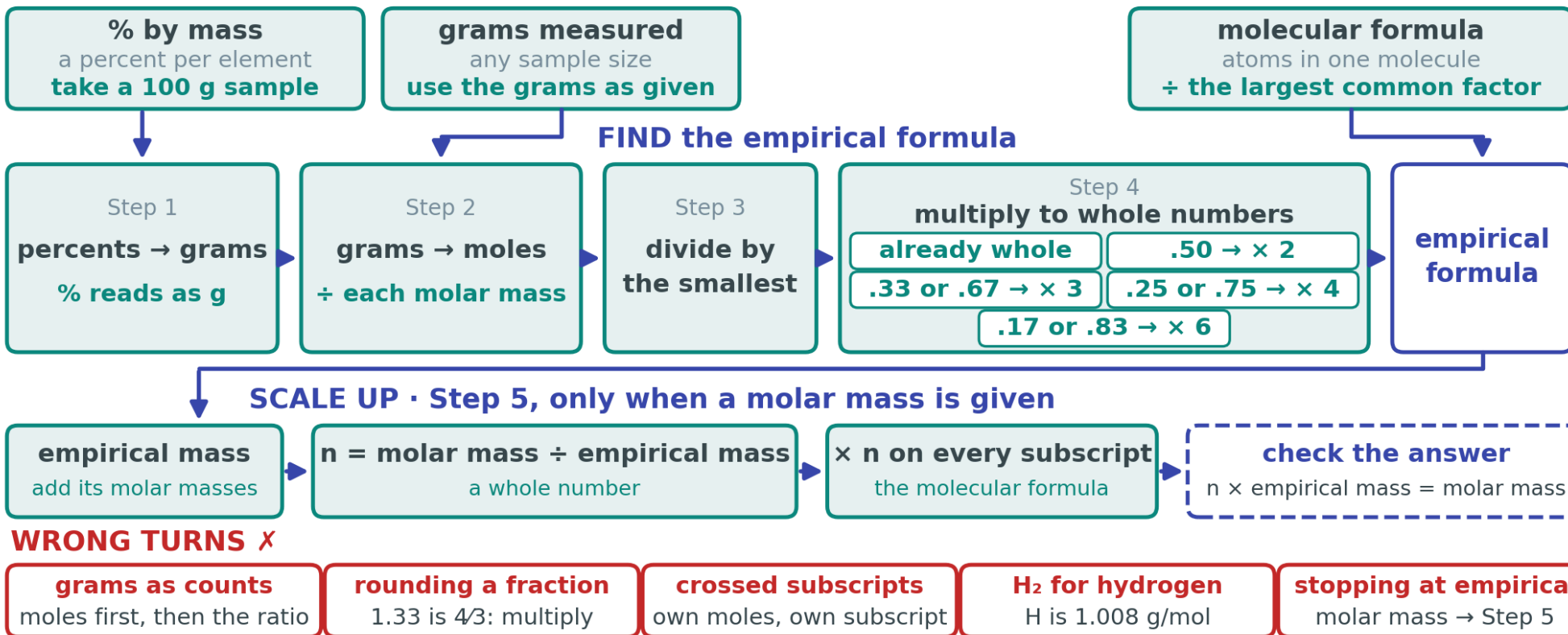


1. **Percents → grams:** take a 100 g sample.
2. **Grams → moles:** each element's molar mass.
3. **Divide by the smallest** mole count.
4. **Multiply to whole numbers.**
5. **Scale up:** $n = \text{molar mass} \div \text{empirical mass}$.

Percent to mass · mass to mole · divide by small · multiply 'til whole

One route for every formula problem

START from the data



Start from the data. Steps 1 to 4 reach the empirical formula; a sixth, .17 or .83, takes × 6. Step 5 needs a given molar mass. Wrong turns sit off the route.



Guided example: a sulfur oxide

sulfur oxide: 40.05% S, 59.95% O by mass

given: mass percents · wanted: empirical formula

Coal smoke carries a sulfur oxide that is **40.05% S** and **59.95% O** by mass. Find its empirical formula. (*S 32.07 · O 16.00 g/mol*)

Name each move as it happens: percents, grams, moles, ratio.

✓ Guided example: grams, then moles

sulfur oxide: 40.05% S, 59.95% O by mass

given: mass percents · wanted: empirical formula

Step 1 · Percents → grams

Step 2 · Grams → moles

$$40.05 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 1.249 \text{ mol S} \cdot 59.95 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.747 \text{ mol O}$$

🧠 **Sense check:** Oxygen has more grams and lighter atoms, so it has far more moles: 3.747 against 1.249. ✓

✓ Guided example: the smallest ratio

sulfur oxide: 40.05% S, 59.95% O by mass
moles, per 100 g: S 1.249 · O 3.747

Step 3 · Divide by the smallest

$$\text{S: } \frac{1.249 \text{ mol}}{1.249} = 1.00 \cdot \text{O: } \frac{3.747 \text{ mol}}{1.249} = 3.00$$

Step 4 · Multiply to whole numbers

empirical formula: SO_3

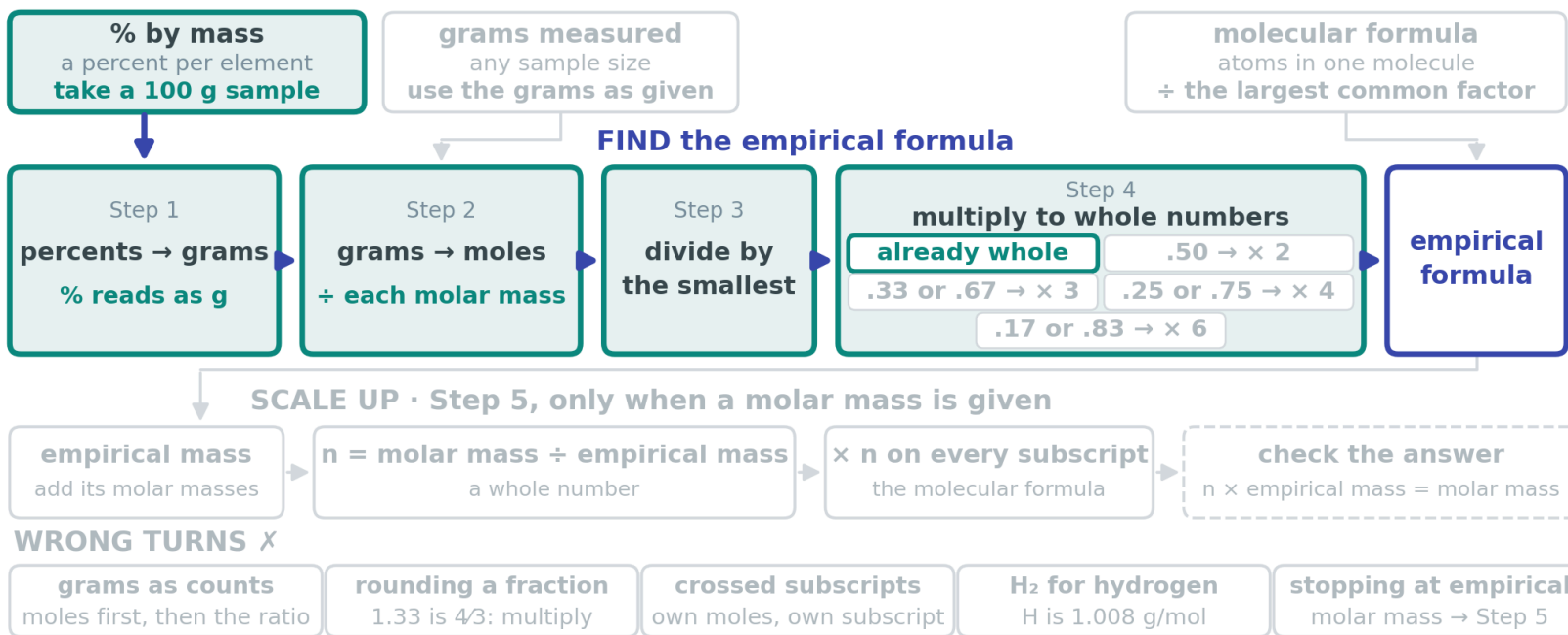
1 S : 3 O · the smallest whole-number atom ratio

🧠 **Sense check:** Rebuild the data: sulfur is 32.07 of SO_3 's 80.07 g/mol, which is 40.05%: the given percent returns. ✓

✓ Guided example: the route on the map

sulfur oxide: 40.05% S, 59.95% O by mass
found: SO_3

START from the data



🧠 **Sense check:** Percents given: Steps 1 to 4. The ratio landed whole, and no molar mass was given. ✓



Practice 1

hexane: C_6H_{14}

given: molecular formula · wanted: empirical formula

Hexane, C_6H_{14} , extracts cooking oil from soybeans. What is its empirical formula?

- A. C_3H_7
- B. CH_2
- C. C_6H_{14}
- D. CH

✓ Practice 1: answer A

hexane: C_6H_{14}

wanted: empirical formula



6 and 14 share a factor of 2 · 3 and 7 share none

B divided by the smallest subscript, 6: $14 \div 6 = 2.33$, then rounded the fraction away to CH_2 . C kept the molecular formula: 6 and 14 still share a factor of 2. D set every subscript to 1: the empirical formula keeps the ratio, not one of each.

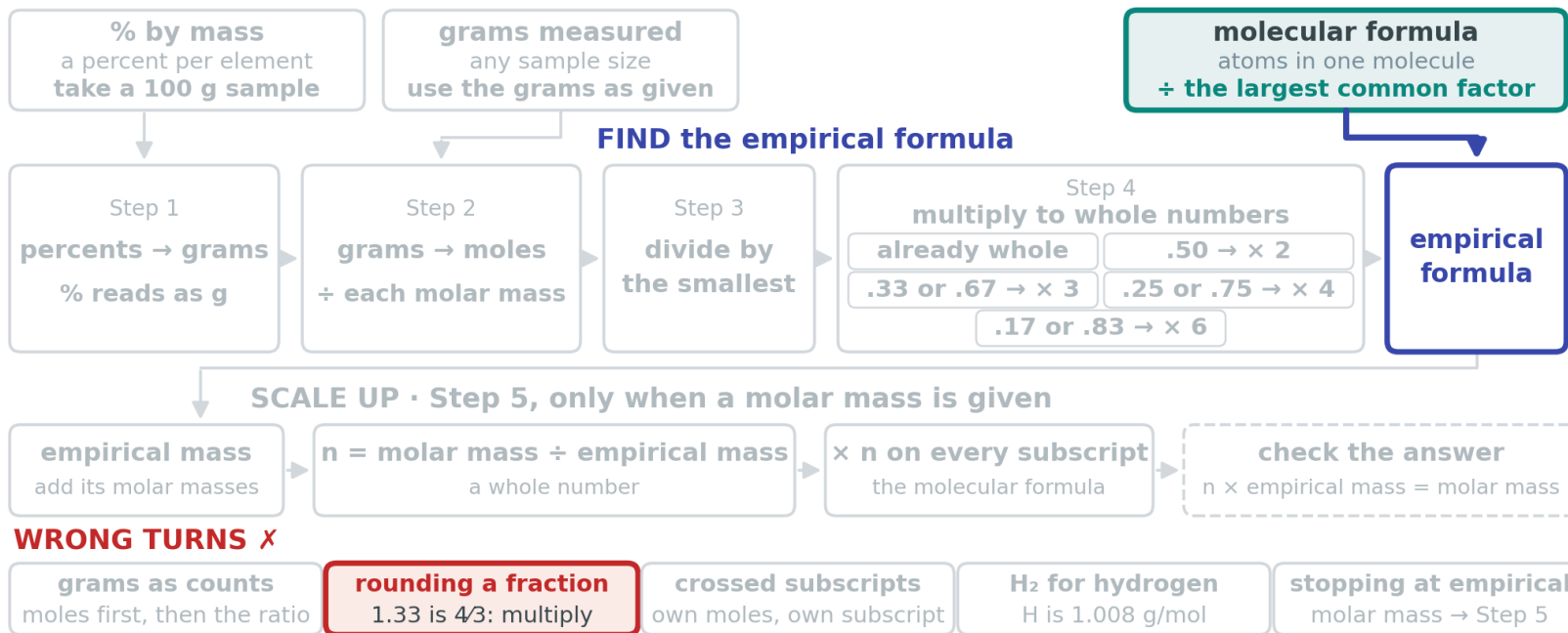
🧠 **Sense check:** C_3H_7 taken twice rebuilds C_6H_{14} , and 3 : 7 reduces no further. ✓

✓ Practice 1: the route on the map

hexane: C_6H_{14}

found: C_3H_7

START from the data



🧠 **Sense check:** A molecular formula given outright skips Steps 1 to 4: divide by the largest common factor. ✓



Worked example 1

unknown white solid: 40.0% C, 6.7% H, 53.3% O by mass

given: mass percents · wanted: empirical formula

A white solid arrives for analysis: **40.0% C**, **6.7% H**, **53.3% O** by mass. Find its empirical formula. (*C 12.01 · H 1.008 · O 16.00 g/mol*)

Write the route first: percents → grams → moles → smallest whole-number ratio.

✓ Worked example 1: grams, then moles

unknown white solid: 40.0% C, 6.7% H, 53.3% O by mass

given: mass percents · wanted: empirical formula

Step 1 · Percents → grams

Step 2 · Grams → moles

$$40.0 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.33 \text{ mol C} \cdot 6.7 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 6.65 \text{ mol H}$$

$$53.3 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.33 \text{ mol O}$$

 **Sense check:** Hydrogen has the smallest mass, 6.7 g, yet the most moles: hydrogen atoms are the lightest. Mass order is not mole order. ✓

✓ Worked example 1: the smallest ratio

unknown white solid: 40.0% C, 6.7% H, 53.3% O by mass

moles, per 100 g: C 3.33 · H 6.65 · O 3.33

Step 3 · Divide by the smallest

$$\text{C: } \frac{3.33 \text{ mol}}{3.33} = 1.00 \cdot \text{H: } \frac{6.65 \text{ mol}}{3.33} = 2.00 \cdot \text{O: } \frac{3.33 \text{ mol}}{3.33} = 1.00$$

Step 4 · Multiply to whole numbers

empirical formula: CH_2O

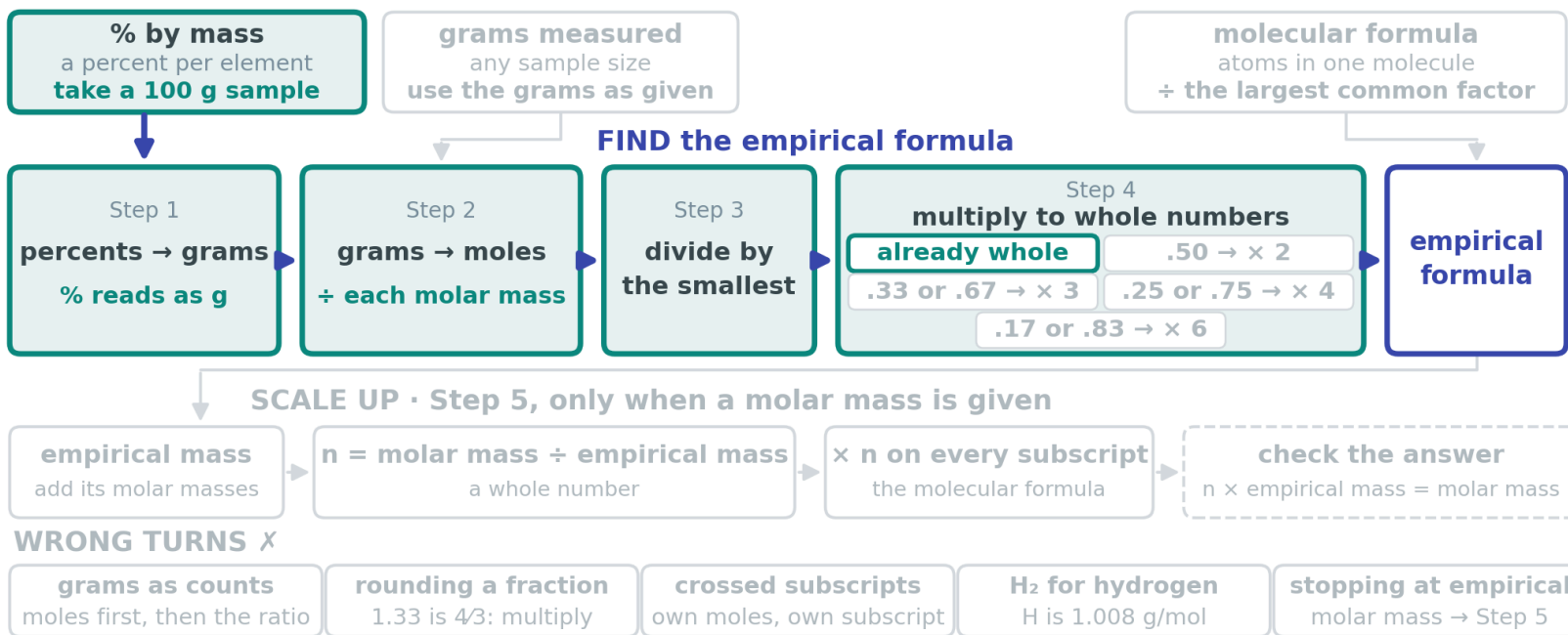
1 C : 2 H : 1 O · the smallest whole-number atom ratio

🧠 **Sense check:** Rebuild the data from the formula: carbon is 12.01 of CH_2O 's 30.03 g/mol, which is 40.0%: the given percent returns. ✓

✓ Worked example 1: the route on the map

unknown white solid: 40.0% C, 6.7% H, 53.3% O by mass
found: CH₂O

START from the data



🧠 **Sense check:** The sulfur oxide path: Steps 1 to 4, an already-whole ratio. A third element adds one conversion. ✓



Worked example 2: magnetite

magnetite: 72.4% Fe, 27.6% O by mass

given: mass percents · wanted: empirical formula

Magnetite is the naturally magnetic iron ore. It is **72.4% Fe** and **27.6% O** by mass. Find the empirical formula. (*Fe 55.85 · O 16.00 g/mol*)

Convert the percents to moles and divide by the smallest.

✓ Worked example 2: grams, then moles

magnetite: 72.4% Fe, 27.6% O by mass
given: mass percents · wanted: empirical formula

Step 1 · Percents → grams

Step 2 · Grams → moles

$$72.4 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 1.296 \text{ mol Fe} \cdot 27.6 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.725 \text{ mol O}$$

🧠 **Sense check:** Iron carries most of the mass yet has fewer moles: one Fe atom outweighs three O atoms. ✓

✓ Worked example 2: clearing the fraction

magnetite: 72.4% Fe, 27.6% O by mass

Step 3 • Divide by the smallest

$$\text{Fe: } \frac{1.296 \text{ mol}}{1.296} = 1.00 \cdot \text{O: } \frac{1.725 \text{ mol}}{1.296} = 1.33$$

FeO → 77.7% Fe, but the sample is 72.4% Fe ✗
test the rounded formula against the data

Step 4 • Multiply to whole numbers

$$\text{Fe: } 1.00 \times 3 = 3.00 \cdot \text{O: } 1.33 \times 3 = 3.99 \rightarrow 4 = \text{Fe}_3\text{O}_4$$

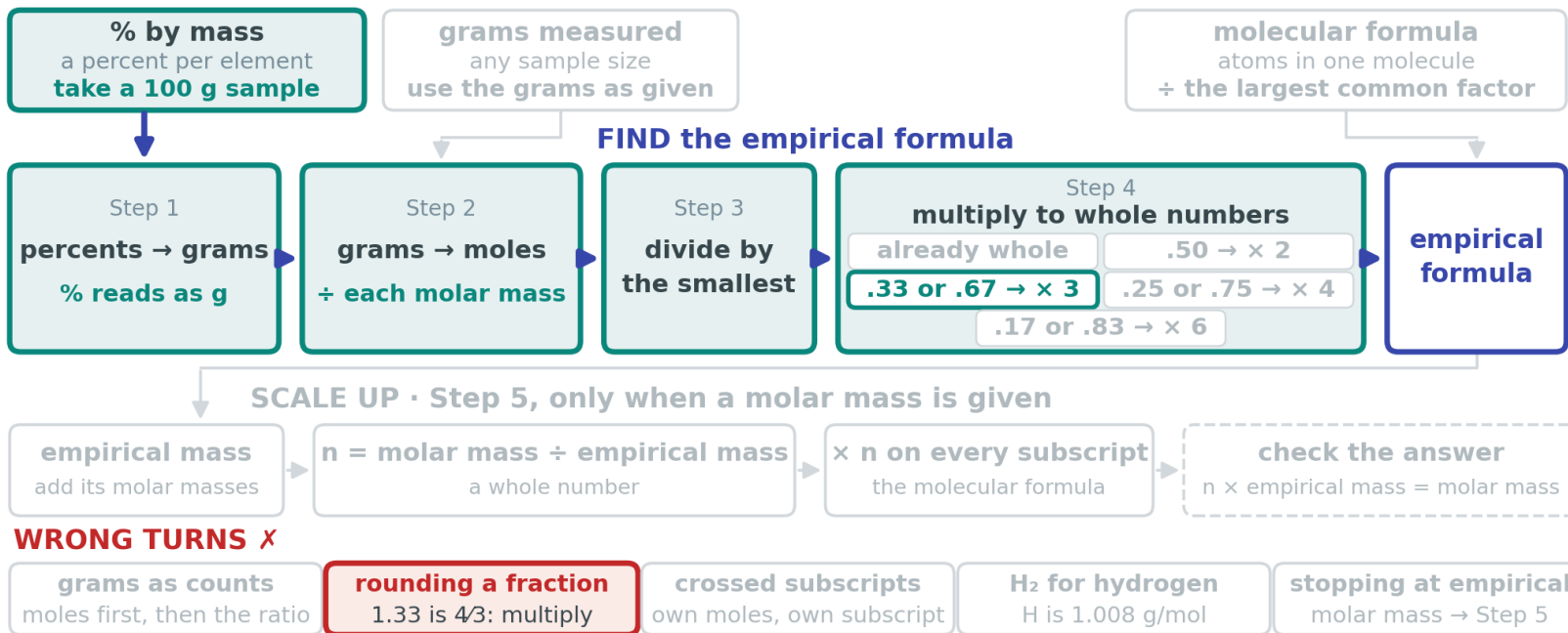
 **Sense check:** Fe₃O₄ rebuilds to 72.4% Fe: the data agree. ✓

✓ Worked example 2: the route on the map

magnetite: 72.4% Fe, 27.6% O by mass

tested: FeO ✗ · found: Fe₃O₄ ✓

START from the data



🧠 **Sense check:** Rounding 1.33 to 1 is a wrong turn: FeO misses the data. Step 4 takes the × 3 chip instead. ✓

Take-home: clear fractions by multiplying

$$1 : 1.33 = 1 : 4/3 \rightarrow \times 3 \rightarrow 3 : 4 \rightarrow \text{Fe}_3\text{O}_4$$

rounded instead to 1 : 1 $\rightarrow$ FeO, 77.7% Fe against 72.4% measured $\times$

$$.50 \rightarrow \times 2 \cdot .33 \text{ or } .67 \rightarrow \times 3 \cdot .25 \text{ or } .75 \rightarrow \times 4 \cdot .17 \text{ or } .83 \rightarrow \times 6$$

multiply every element's count by the same factor $\cdot$ round only near-whole results such as 2.99 $\rightarrow$ 3

A ratio that lands on a clean fraction is exact: rounding it away changes the compound.
Multiply every mole count by the fraction's denominator.

Your turn: a phosphorus oxide

phosphorus oxide: 43.6% P, 56.4% O by mass

given: mass percents · wanted: empirical formula

$$43.6 \text{ g P} \times \frac{1 \text{ mol P}}{30.97 \text{ g P}} = 1.41 \text{ mol P} \cdot 56.4 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.53 \text{ mol O}$$

$$\text{P: } \frac{1.41 \text{ mol}}{1.41} = 1.00 \cdot \text{O: } \frac{3.53 \text{ mol}}{1.41} = \underline{\hspace{1cm}} \cdot \text{multiply by } \underline{\hspace{1cm}} = \underline{\hspace{1cm}}$$

$$\text{P: } 1.00 \cdot \text{O: } \frac{3.53 \text{ mol}}{1.41} = 2.50 \cdot \times 2 \rightarrow \text{P } 2.00, \text{O } 5.00 = \mathbf{P_2O_5}$$



Practice 2

aluminum oxide coat: 52.9% Al, 47.1% O by mass

given: mass percents · wanted: empirical formula

Bare aluminum forms a thin oxide coat within seconds in air. The coat is **52.9% Al** and **47.1% O** by mass. What is its empirical formula? (Al 26.98 · O 16.00 g/mol)

- A. AlO
- B. Al₂O₃
- C. AlO₂
- D. Al₃O₂

✓ Practice 2: answer B

aluminum oxide coat: 52.9% Al, 47.1% O by mass

given: mass percents · wanted: empirical formula

$$52.9 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 1.96 \text{ mol Al} \cdot 47.1 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.94 \text{ mol O}$$

$$\text{Al: } \frac{1.96 \text{ mol}}{1.96} = 1.00 \cdot \text{O: } \frac{2.94 \text{ mol}}{1.96} = 1.50 \cdot \times 2 \rightarrow \text{Al } 2, \text{O } 3 = \mathbf{Al_2O_3 \text{ (answer B)}}$$

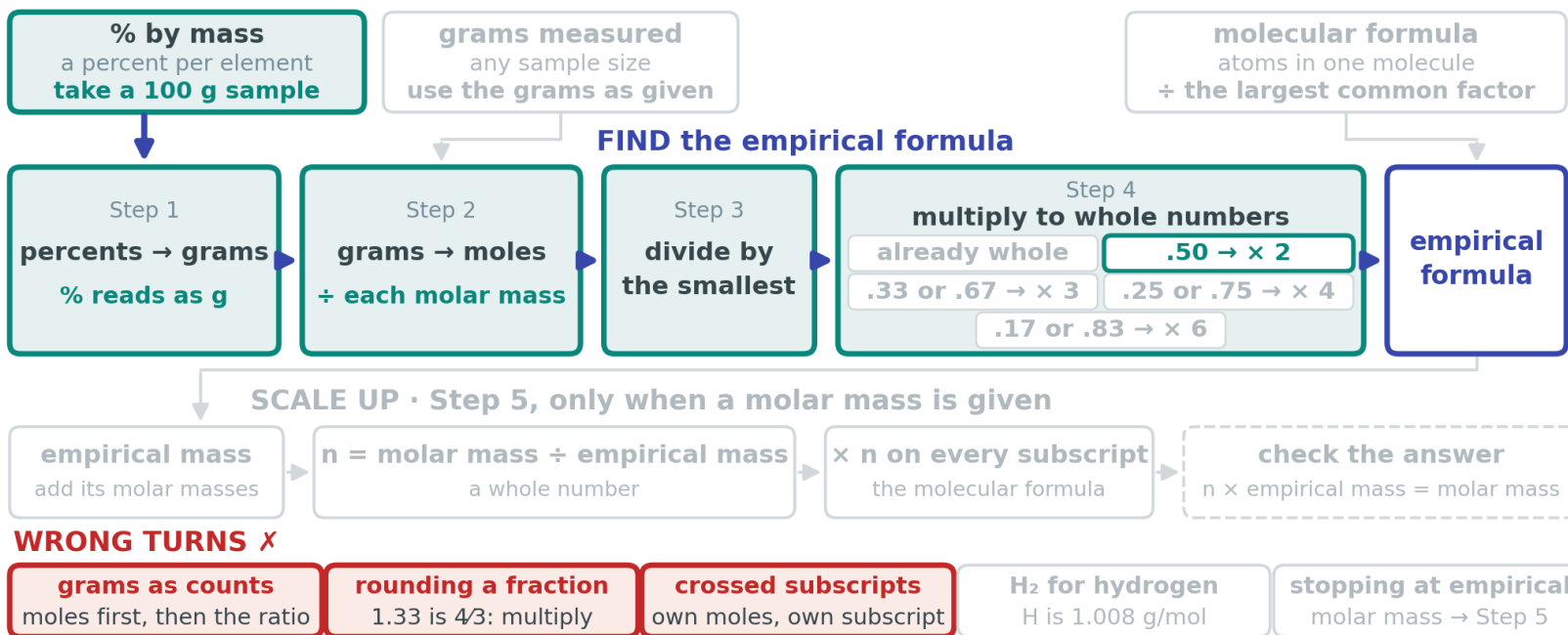
A compared the grams directly: $52.9 \div 47.1 = 1.12$ reads as 1 : 1, but grams are not counts. C rounded 1.50 up to 2; 1.50 is $\frac{3}{2}$, and fractions multiply away. D crossed the mole amounts: O's 2.94 mol is the larger count, so O takes the larger subscript.

 **Sense check:** Al_2O_3 rebuilds to 52.9% Al ✓; oxygen is less than half the mass but three atoms of every five.

✓ Practice 2: the route on the map

aluminum oxide coat: 52.9% Al, 47.1% O by mass
found: Al_2O_3

START from the data



🧠 **Sense check:** Steps 1 to 4 with the × 2 chip: 1.50 is $\frac{3}{2}$. A, C and D each took a red wrong turn. ✓

The multiplier n

$$n = \text{molar mass} \div \text{empirical formula mass}$$

$$n = 180.16 \div 30.03 = 6 \rightarrow \text{CH}_2\text{O} \times 6 = \text{C}_6\text{H}_{12}\text{O}_6$$

Every multiple of CH_2O has the same percent composition, so percents alone stop at the empirical formula. The measured molar mass picks out the multiple: the molecular formula is the empirical unit taken n times.



Worked example 3: the molecular formula

unknown white solid: 40.0% C, 6.7% H, 53.3% O · molar mass 180.16 g/mol

given: mass percents and molar mass · wanted: molecular formula

A white solid is **40.0% C**, **6.7% H**, **53.3% O** by mass; a separate measurement gives its molar mass, **180.16 g/mol**. Find the molecular formula. (C 12.01 · H 1.008 · O 16.00 g/mol)

Find the empirical formula first, then scale up to the molar mass.

✓ Worked example 3: scale to the molar mass

unknown solid: 40.0% C, 6.7% H, 53.3% O · 180.16 g/mol

given: mass percents and molar mass · wanted: molecular formula

Steps 1 to 4 · Reach the empirical formula

empirical formula CH_2O : $12.01 + 2(1.008) + 16.00 = 30.03 \text{ g/mol}$

Scale up · $n = \text{molar mass} \div \text{empirical mass}$

$$n = \frac{180.16 \text{ g/mol}}{30.03 \text{ g/mol}} = 6.00 \cdot \text{CH}_2\text{O} \times 6 = \text{C}_6\text{H}_{12}\text{O}_6$$

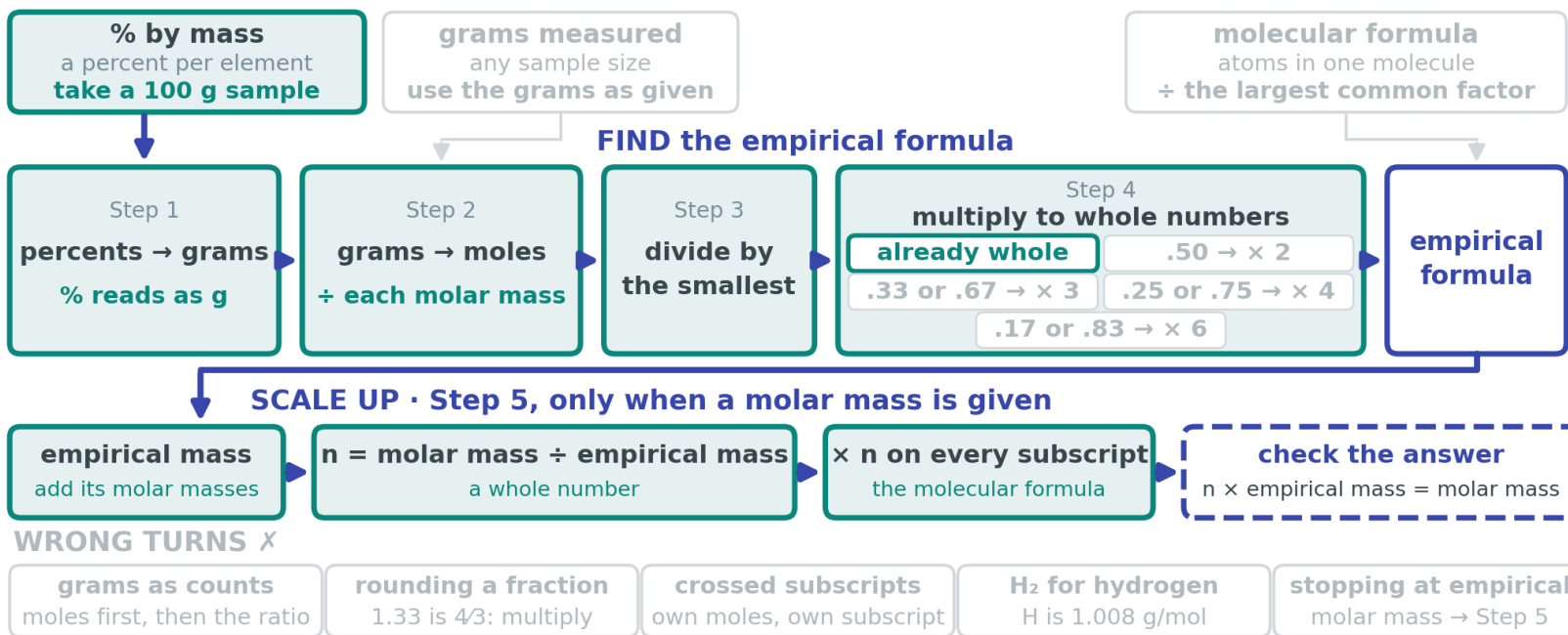
 **Sense check:** n landed whole, and $6 \times 30.03 = 180.2$ rebuilds the molar mass. An unknown white powder with these numbers is glucose, blood sugar. ✓

✓ Worked example 3: the route on the map

unknown solid: 40.0% C, 6.7% H, 53.3% O · 180.16 g/mol

found: CH₂O, then C₆H₁₂O₆

START from the data



🧠 **Sense check:** The CH₂O route, then Step 5: a molar mass was given, so the empirical formula scales by n = 6. ✓

⚠ Where this goes wrong

Rounding the fraction away. Magnetite's mole ratio, $1.725 \div 1.296 = 1.33$, rounded to 1 : 1 gives FeO: 77.7% Fe against the measured 72.4%. 1.33 is $4/3$: multiply every count by 3 $\rightarrow$ Fe₃O₄.

Stopping at the empirical formula. 40.0% C, 6.7% H, 53.3% O gives CH₂O, 30.03 g/mol. A measured molar mass of 180.16 g/mol demands $n = 6$: the molecular formula is C₆H₁₂O₆.

Scaling without checking n. C₁₂H₂₄O₁₂ weighs $12 \times 30.03 = 360.4$ g/mol, double the 180.16 given. Upside down, $30.03 \div 180.16 = 0.167$, and no molecule holds a sixth of a unit. $n \times$ empirical mass must rebuild the molar mass.

Crossing the mole amounts. Magnetite holds 1.296 mol Fe and 1.725 mol O per 100 g. The larger count belongs to O, so O takes the larger subscript: Fe₃O₄, never Fe₄O₃. Each subscript comes from its own element's moles.

 Practice 3

hydrocarbon vapor: 92.3% C, 7.7% H · molar mass 78.11 g/mol

given: mass percents and molar mass · wanted: molecular formula

An industrial solvent's vapor is **92.3% C** and **7.7% H** by mass, with molar mass **78.11 g/mol**. Which molecular formula is consistent with these data? (C 12.01 · H 1.008 g/mol)

- A. CH
- B. C₆H₃
- C. C₆H₆
- D. C₆H₁₂

✓ Practice 3: answer C

hydrocarbon vapor: 92.3% C, 7.7% H · molar mass 78.11 g/mol

$$92.3 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 7.69 \text{ mol C} \cdot 7.7 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 7.64 \text{ mol H}$$

$$\text{C: } \frac{7.69}{7.64} = 1.01 \cdot \text{H: } \frac{7.64}{7.64} = 1.00 \rightarrow \text{CH (13.02)} \cdot n = \frac{78.11}{13.02} = 6 = \text{C}_6\text{H}_6 \text{ (answer C)}$$

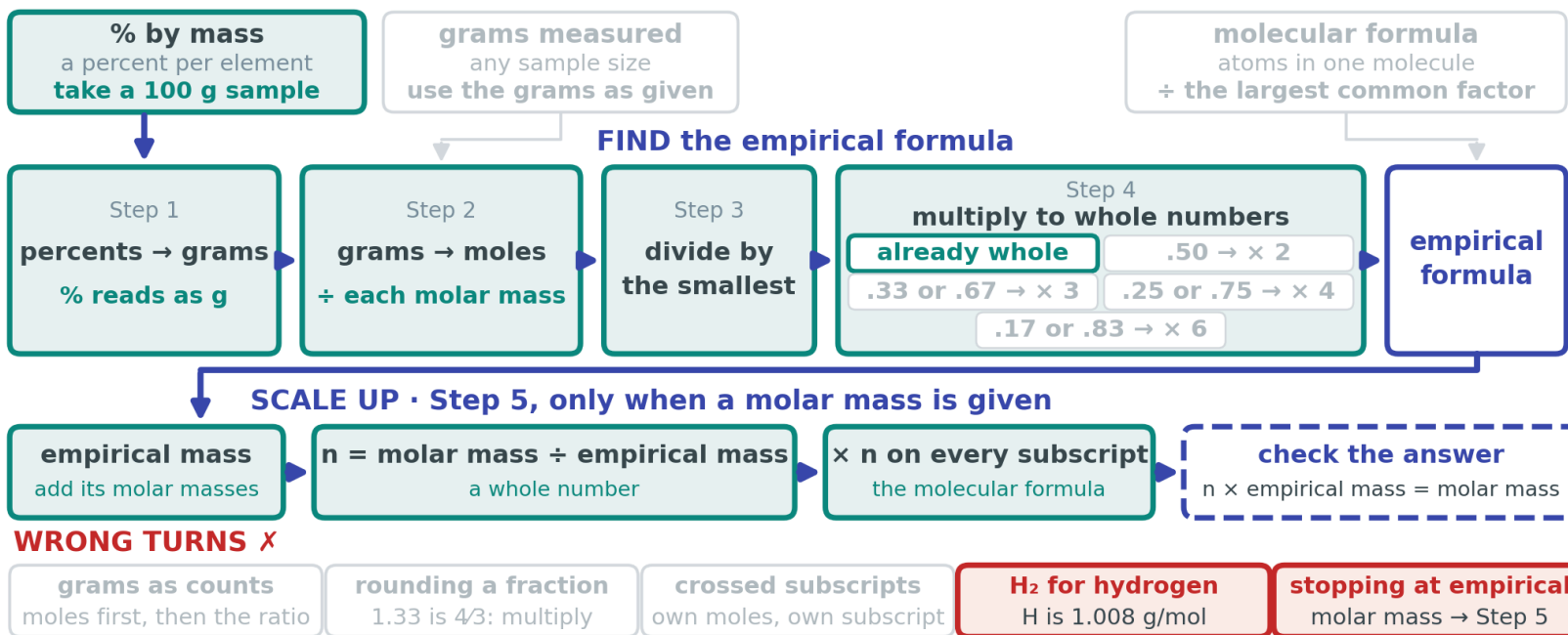
A stopped at CH: 13.02 g/mol, not 78.11. B used H₂'s 2.016 for hydrogen: 7.7 ÷ 2.016 = 3.82 mol → C₂H, n = 78.11 ÷ 25.03 = 3.12 → C₆H₃. D doubled the H moles: 2 × 7.64 = 15.3 → CH₂, n = 78.11 ÷ 14.03 = 5.57 → C₆H₁₂. Neither wrong n landed whole.

🧠 **Sense check:** n = 6 lands whole: 6 × 13.02 = 78.1. B (75.08) and D (84.16) sit close: mass alone cannot decide. ✓

✓ Practice 3: the route on the map

hydrocarbon vapor: 92.3% C, 7.7% H · 78.11 g/mol
found: CH, then C₆H₆

START from the data



🧠 **Sense check:** The glucose route, then Step 5 with n = 6. A stopped early; B took H₂. ✓



Practice 4

vitamin C: 40.91% C, 4.58% H, 54.51% O by mass · molar mass 176.12 g/mol
given: mass percents and molar mass · wanted: molecular formula

Vitamin C is **40.91% C**, **4.58% H** and **54.51% O** by mass; one mole weighs **176.12 g**.
What is the molecular formula of vitamin C? (C 12.01 · H 1.008 · O 16.00 g/mol)

- A. $\text{C}_3\text{H}_4\text{O}_3$
- B. $\text{C}_6\text{H}_6\text{O}_6$
- C. $\text{C}_6\text{H}_4\text{O}_6$
- D. $\text{C}_6\text{H}_8\text{O}_6$

✓ Practice 4: answer D

vitamin C: 40.91% C, 4.58% H, 54.51% O · 176.12 g/mol

moles, per 100 g: C 3.406 · H 4.54 · O 3.407

$$\text{C: } \frac{3.406}{3.406} = 1.00 \cdot \text{H: } \frac{4.54}{3.406} = 1.33 \cdot \text{O: } \frac{3.407}{3.406} = 1.00 \cdot \times 3 \rightarrow \text{C}_3\text{H}_4\text{O}_3 (88.06 \text{ g/mol})$$

$$n = \frac{176.12 \text{ g/mol}}{88.06 \text{ g/mol}} = 2.000 \cdot \text{C}_3\text{H}_4\text{O}_3 \times 2 = \text{C}_6\text{H}_8\text{O}_6 \text{ (answer D)}$$

A stopped at the empirical formula, 88.06 g/mol: half the molar mass. B rounded 1.33 to 1: CHO, $n = 176.12 \div 29.02 = 6.07$, rounded to 6. C used H₂'s 2.016: $4.58 \div 2.016 = 2.27$ mol H, C₃H₂O₃, $n = 176.12 \div 86.05 = 2.05$.

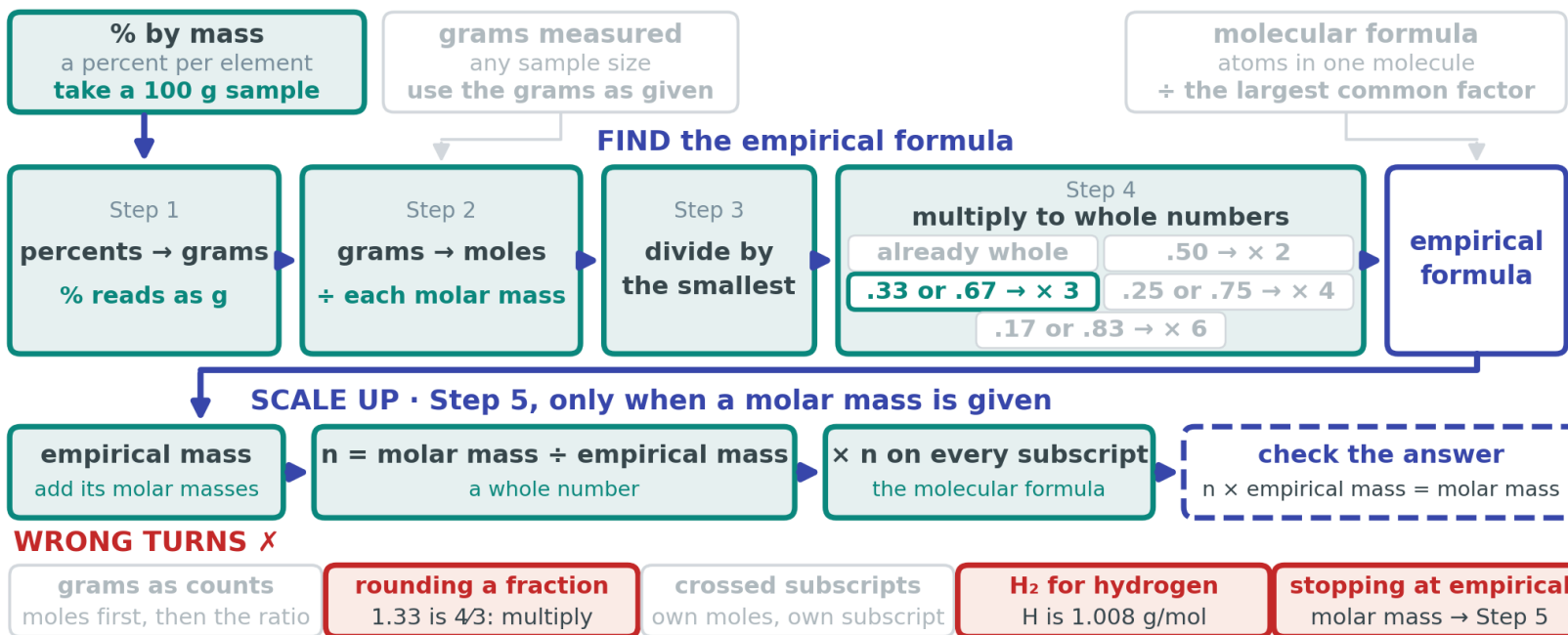
🧠 **Sense check:** Only $n = 2.000$ lands whole: $2 \times 88.06 = 176.12$ ✓

✓ Practice 4: the route on the map

vitamin C: 40.91% C, 4.58% H, 54.51% O · 176.12 g/mol

found: $C_3H_4O_3$, then $C_6H_8O_6$

START from the data



🧠 **Sense check:** The × 3 chip, then Step 5 with n = 2. Each wrong choice took one red turn. ✓



Extra practice: a measured sample

hydrocarbon solvent: 1.810 g C and 0.190 g H in a 2.000 g sample

given: measured grams · wanted: empirical formula

An elemental analysis finds **1.810 g C** and **0.190 g H** in a **2.000 g** sample of a hydrocarbon solvent. What is its empirical formula? (C 12.01 · H 1.008 g/mol)

- A. CH
- B. C₄H₅
- C. C₂H₃
- D. C₁₀H
- E. C₅H₄

✓ Extra practice: answer B

hydrocarbon solvent: 1.810 g C and 0.190 g H in a 2.000 g sample

grams measured: no 100 g step

$$1.810 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.1507 \text{ mol C} \cdot 0.190 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.188 \text{ mol H}$$

$$\text{C: } \frac{0.1507}{0.1507} = 1.00 \cdot \text{H: } \frac{0.188}{0.1507} = 1.25 \cdot \times 4 \rightarrow \text{C } 4, \text{H } 5 = \text{C}_4\text{H}_5 \text{ (answer B)}$$

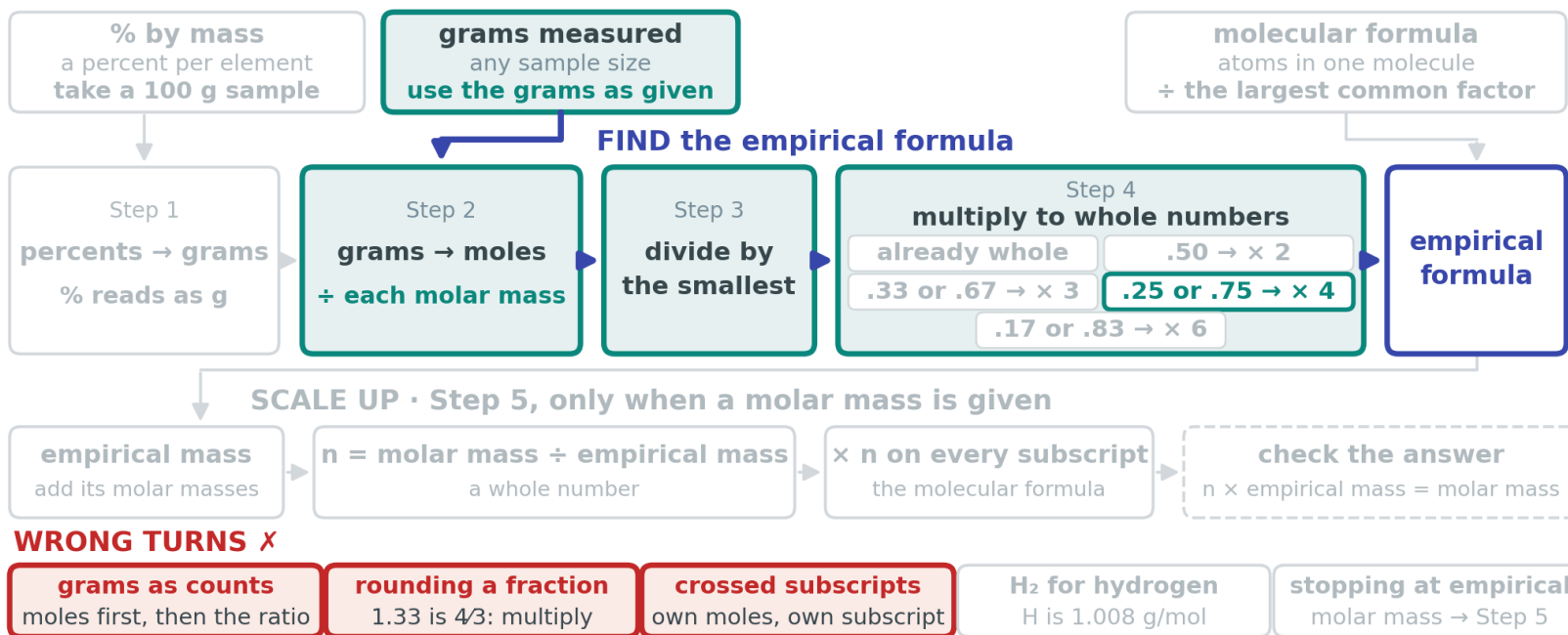
A rounded 1.25 down to 1: CH. C multiplied by 2: $1.25 \times 2 = 2.50$, then rounded up to C_2H_3 . D read grams as counts: $1.810 \div 0.190 = 9.53$, about 10 C per H. E crossed the subscripts: C took hydrogen's 5.

 **Sense check:** Carbon is 48.04 of C_4H_5 's 53.08 g/mol, 90.50%, and the sample is $1.810 \div 2.000 = 90.50\%$ carbon. ✓

✓ Extra practice: the route on the map

hydrocarbon solvent: 1.810 g C and 0.190 g H in a 2.000 g sample
found: C_4H_5

START from the data



🧠 **Sense check:** Grams measured on a 2.000 g sample: skip Step 1 and start at Step 2. The × 4 chip clears the .25. ✓

Check yourself

1. A compound's percent composition is known. State the steps that lead to its empirical formula. (*Where does the 100 g sample enter, and where do moles enter?*)
2. Dividing by the smallest mole count leaves a ratio of 1 : 2.50. State the next move and the whole numbers it produces.

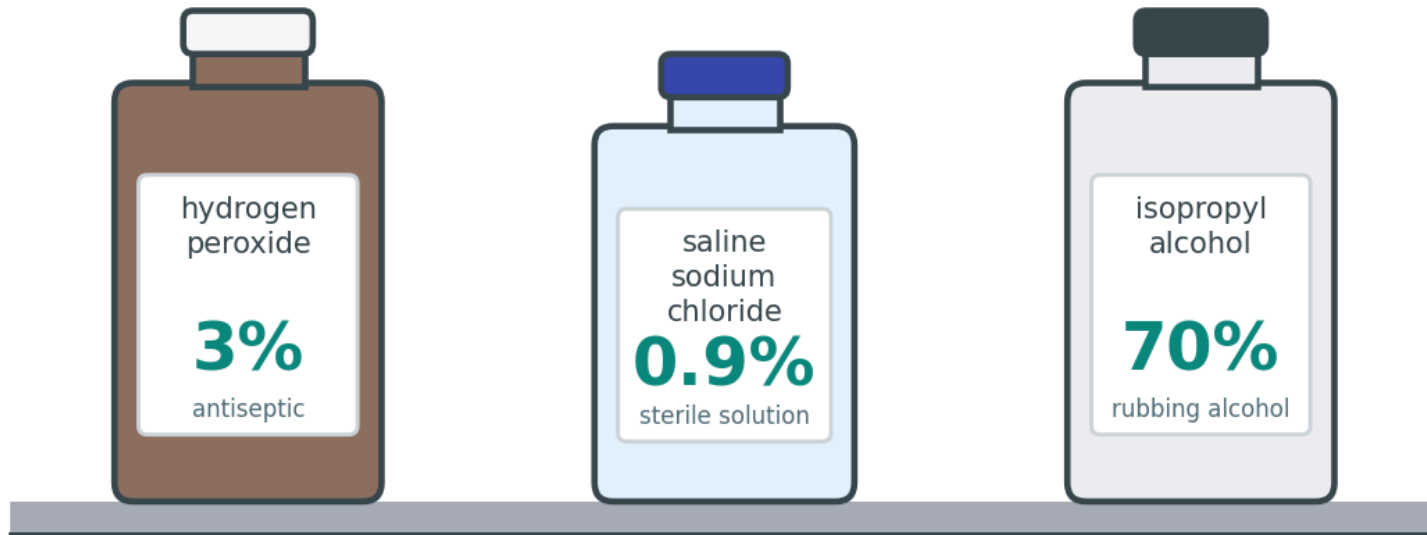
Percent also states a solution's strength: 5.0% by mass means 5.0 g of solute in every 100 g of solution: the same parts-per-hundred idea, used as a conversion factor.

4 • Percent Concentration

Compute a solution's mass, volume, or mass-volume percent, and use a labeled percent as a conversion factor between the amount of solution and the amount of solute.

Three labels, three percents

one drugstore shelf, three labels, three percents



Peroxide reads 3%. Saline reads 0.9%. Rubbing alcohol reads 70%. On every label, the number compares the active ingredient to everything in the bottle.

Concentration: solute compared to solution

concentration = amount of solute ÷ amount of solution

solute: the substance dissolved · solvent: what it dissolves in, usually water · solution = solute + solvent

More solute alone does not mean more concentrated.

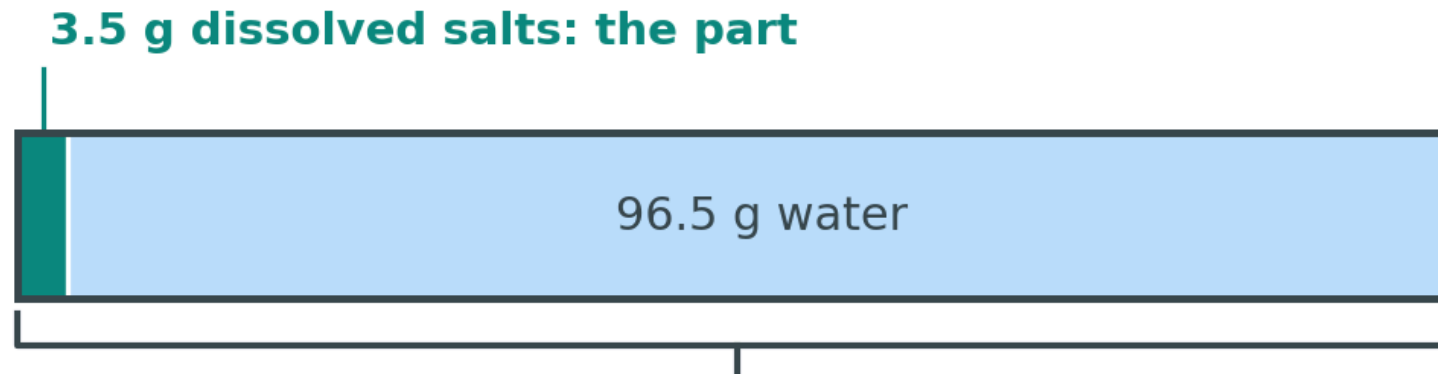
Which drink is more concentrated: 10.0 g of sugar in 200. g of drink, or 15.0 g in 500. g?

$$\frac{10.0 \text{ g sugar}}{200. \text{ g drink}} = 5.00 \text{ g per } 100 \text{ g } \checkmark \quad \frac{15.0 \text{ g sugar}}{500. \text{ g drink}} = 3.00 \text{ g per } 100 \text{ g}$$

The first drink, with less sugar.

Percent concentration: parts per hundred

seawater, 3.5% (m/m): 3.5 parts of salt in every 100 parts of solution



3.5 g + 96.5 g = 100 g of solution: the whole

the denominator is the whole bar, water and salts together

Seawater carries 3.5 g of dissolved salts in every 100 g. A percent concentration states the parts of solute in every hundred parts of solution. The hundred is the whole solution: solute plus solvent.

Mass, volume, and mass-volume percent

one pattern: part over whole solution, times 100; the type sets the units

m/m %	v/v %	m/v %
$\frac{\text{g solute}}{\text{g solution}} \times 100$	$\frac{\text{mL solute}}{\text{mL solution}} \times 100$	$\frac{\text{g solute}}{\text{mL solution}} \times 100$
seawater 3.5 g salts per 100 g	rubbing alcohol 70.0 mL per 100 mL	normal saline 0.90 g NaCl per 100 mL

Each type is the same fraction, part over whole solution, in its own units. Molarity counts moles per liter; a percent needs no molar mass, only a balance or a graduated cylinder.

A percent label is an equality

3.0% (m/m) hydrogen peroxide: 3.0 g H₂O₂ = 100 g solution
every 100 g of solution carries 3.0 g of H₂O₂; the other 97.0 g is water

Every equality gives a conversion factor. This one links solute mass to solution mass:

$$\frac{3.0 \text{ g H}_2\text{O}_2}{100 \text{ g solution}} \text{ or } \frac{100 \text{ g solution}}{3.0 \text{ g H}_2\text{O}_2}$$

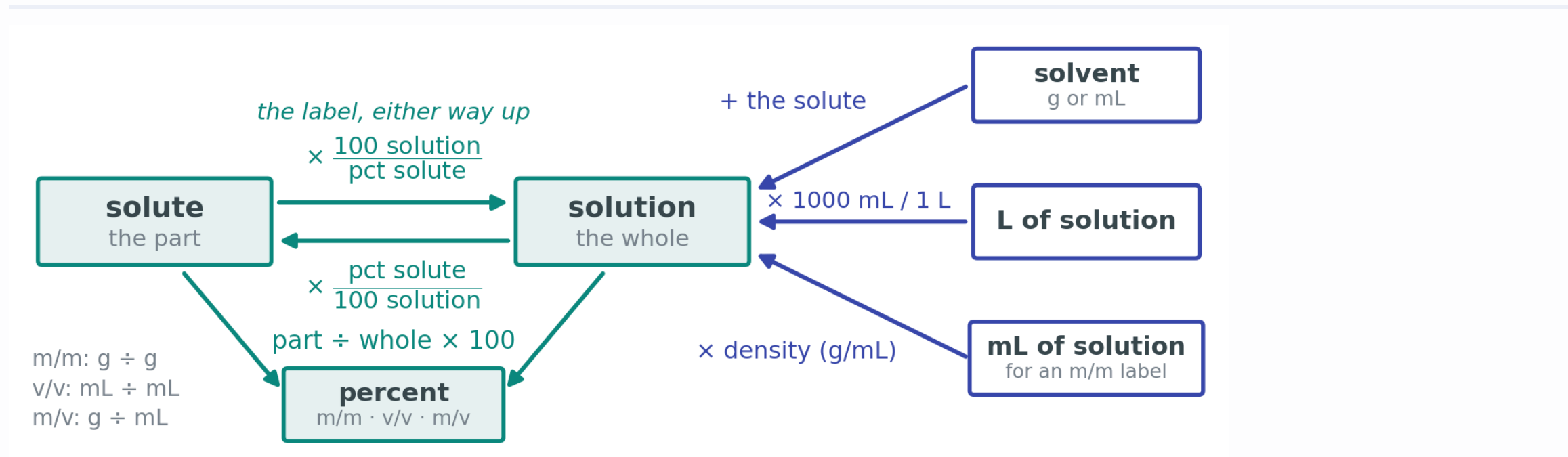
Write it so the **given unit cancels**.



The method

1. **Name part and whole:** the whole solution, solute plus solvent.
2. **Match the units:** m/m g/g, v/v mL/mL, m/v g/mL.
3. **Write the fraction:** part over whole $\times 100$, or the orientation that cancels the given.
4. **Multiply and check.**

One map for every percent problem



To find a percent, divide part by whole, $\times 100$. To use a label, its factor carries solution to solute, or back. Liters, a density, or a separate solvent reach the whole first.



Guided example: sugar in a sports drink

Step 1 • Name part and whole

18.0 g sugar • final volume 0.300 L

part: sugar • whole: 0.300 L of drink, sugar included • wanted: m/v %

A bottle of sports drink carries **18.0 g of sugar** in a final volume of **0.300 L**. What is its mass-volume percent of sugar?

A final volume is already the whole solution. Nothing gets added to it.

✓ Guided example: solution

18.0 g sugar · final volume 0.300 L
part: sugar · whole: 0.300 L of drink · wanted: m/v %

Step 2 · Match the units

$$0.300 \text{ L} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 300. \text{ mL solution}$$

Step 3 · Write the fraction

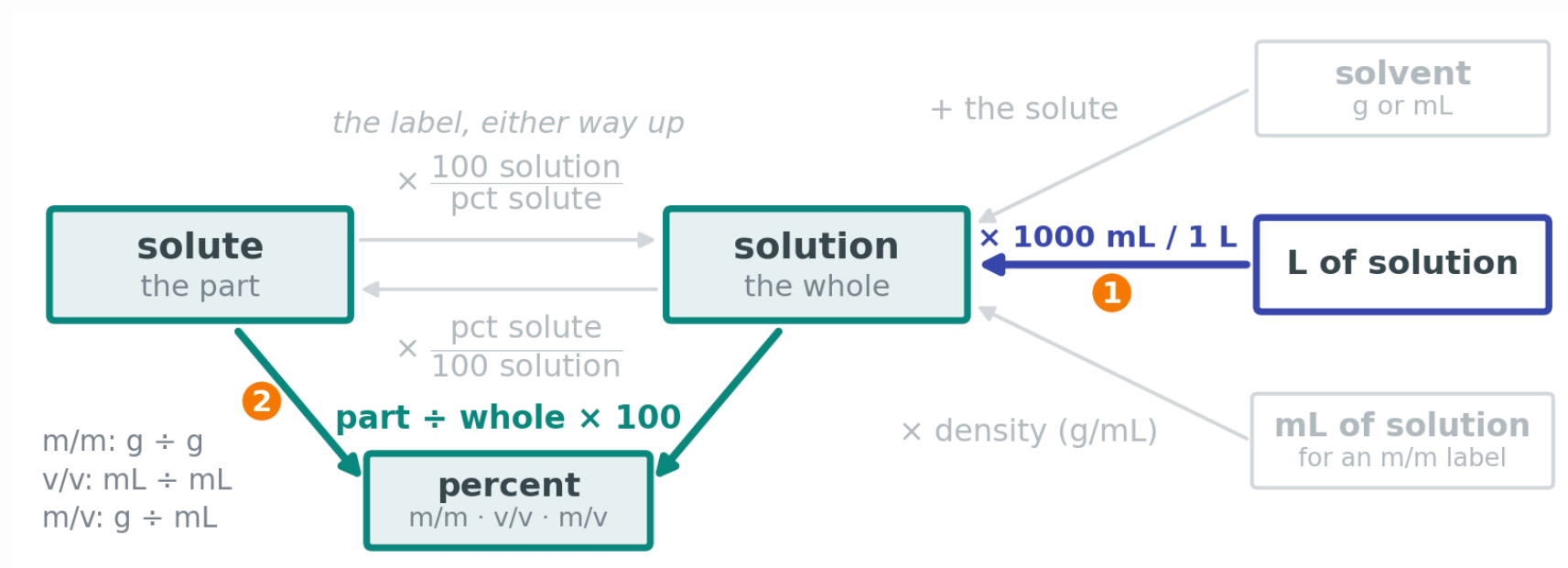
Step 4 · Multiply and check

$$\frac{18.0 \text{ g sugar}}{300. \text{ mL solution}} \times 100 = \mathbf{6.00\% (m/v)}$$

 **Sense check:** 6.00 g of sugar in every 100 mL, and the bottle holds three hundreds: 18.0 g. ✓

✓ Guided example: the route on the map

18.0 g sugar · final volume 0.300 L
given: 18.0 g sugar · 0.300 L of drink · found: 6.00% (m/v)



🧠 **Sense check:** Two moves: liters to milliliters for the whole, then part over whole $\times 100$. ✓



Practice 1

hand sanitizer: 260. mL ethanol in a 400. mL bottle

given: 260. mL solute · 400. mL solution · wanted: v/v %

A **400. mL** bottle of hand sanitizer holds **260. mL** of ethanol. What is the volume percent of ethanol?

- A. 186
- B. 0.650
- C. 65.0
- D. 154
- E. 39.4

✓ Practice 1: answer C

260. mL ethanol in a 400. mL bottle

part: ethanol · whole: 400. mL solution, ethanol included · wanted: v/v %

$$\frac{260. \text{ mL ethanol}}{400. \text{ mL solution}} \times 100 = \mathbf{65.0\% (v/v), \text{ answer C}}$$

A compared the ethanol to the other 140. mL alone: $260. / 140. \times 100 = 186$. B stopped at the decimal fraction: $260. / 400. = 0.650$. D flipped part and whole: $400. / 260. \times 100 = 154$. E added the ethanol to a volume that already held it: $260. / 660. \times 100 = 39.4$.

 **Sense check:** The bottle's 400. mL is the whole, ethanol included. Every 100 mL of sanitizer carries 65.0 mL of ethanol. ✓



Worked example 1: mass percent from masses

25.0 g sucrose dissolved in 100.0 g water

given: 25.0 g solute · 100.0 g solvent · wanted: m/m %

A café batch of simple syrup: **25.0 g of sucrose** dissolved in **100.0 g of water**. What is the mass percent of sucrose?

A common first attempt: divide the 25.0 g of sucrose by the 100.0 g of water. Test the denominator.

✓ Worked example 1: solution

given: 25.0 g sucrose + 100.0 g water · wanted: m/m %

A common first attempt

$$\frac{25.0 \text{ g sucrose}}{100.0 \text{ g water}} \times 100 = 25.0\% \times$$

Step 1 · Name part and whole

Step 2 · Match the units

Step 3 · Write the fraction

Step 4 · Multiply and check

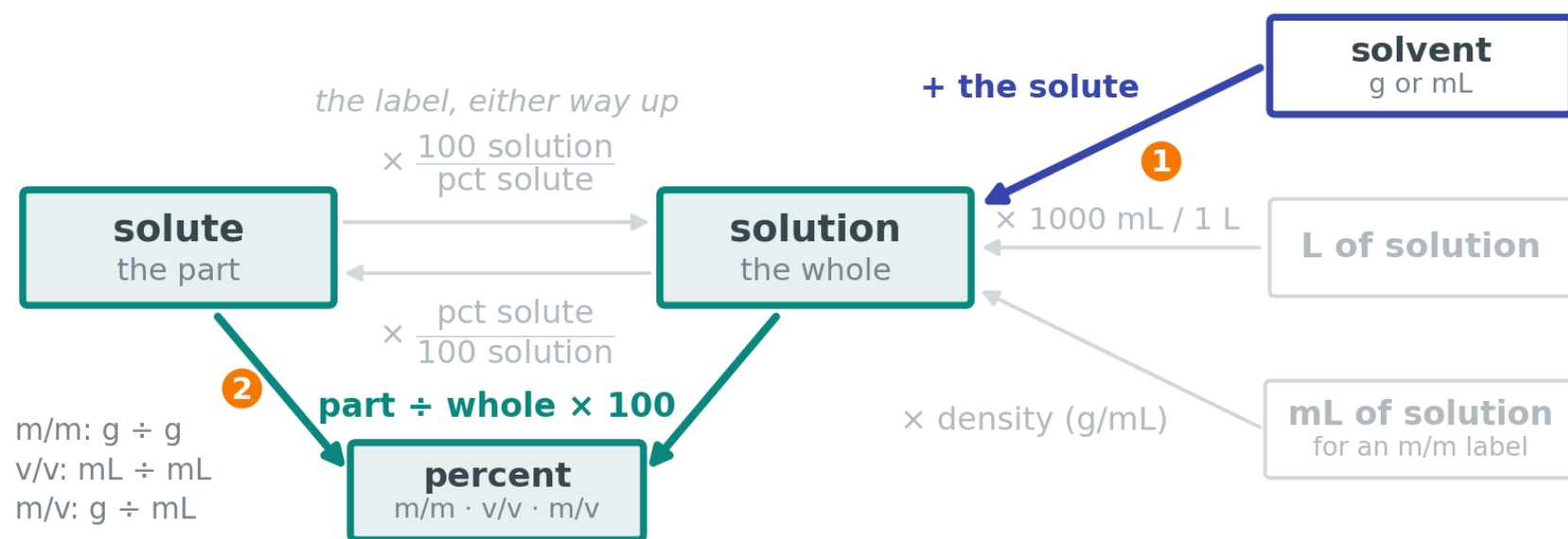
$$\frac{25.0 \text{ g sucrose}}{125.0 \text{ g solution}} \times 100 = \mathbf{20.0\% (m/m)}$$

 **Sense check:** One fifth of 125.0 g is sugar: 20.0 g per 100 g syrup; water alone gives 25.0%, too high. ✓

✓ Worked example 1: the route on the map

25.0 g sucrose dissolved in 100.0 g water

given: 25.0 g solute · 100.0 g solvent · found: 125.0 g solution · 20.0% (m/m)



🧠 **Sense check:** Two moves: the solute joins the solvent to make the whole, then part over whole $\times 100$. ✓

Take-home: the denominator is the whole solution

$$\frac{25.0 \text{ g sucrose}}{100.0 \text{ g water}} \times 100 = 25.0\% \times \text{(solute compared to the solvent alone)}$$

$$\frac{25.0 \text{ g sucrose}}{125.0 \text{ g solution}} \times 100 = \mathbf{20.0\% (m/m) \checkmark}$$

The solute is part of the solution it makes. Add solute and solvent first; that sum is the whole.



Worked example 2: solute mass from the label

Step 1 • Name part and whole

3.0% (m/m) hydrogen peroxide: 3.0 g H_2O_2 = 100 g solution

part: H_2O_2 • whole: solution • given: 250.0 g solution • wanted: g H_2O_2

A drugstore bottle holds **250.0 g of 3.0% (m/m) hydrogen peroxide** solution. What mass of H_2O_2 is in the bottle?

Set it up: which orientation of the percent factor cancels g solution?

✓ Worked example 2: solution

3.0% (m/m) hydrogen peroxide: 3.0 g H₂O₂ = 100 g solution
 part: H₂O₂ · whole: solution · given: 250.0 g solution · wanted: g H₂O₂

Step 2 · Match the units

Step 3 · Write the fraction

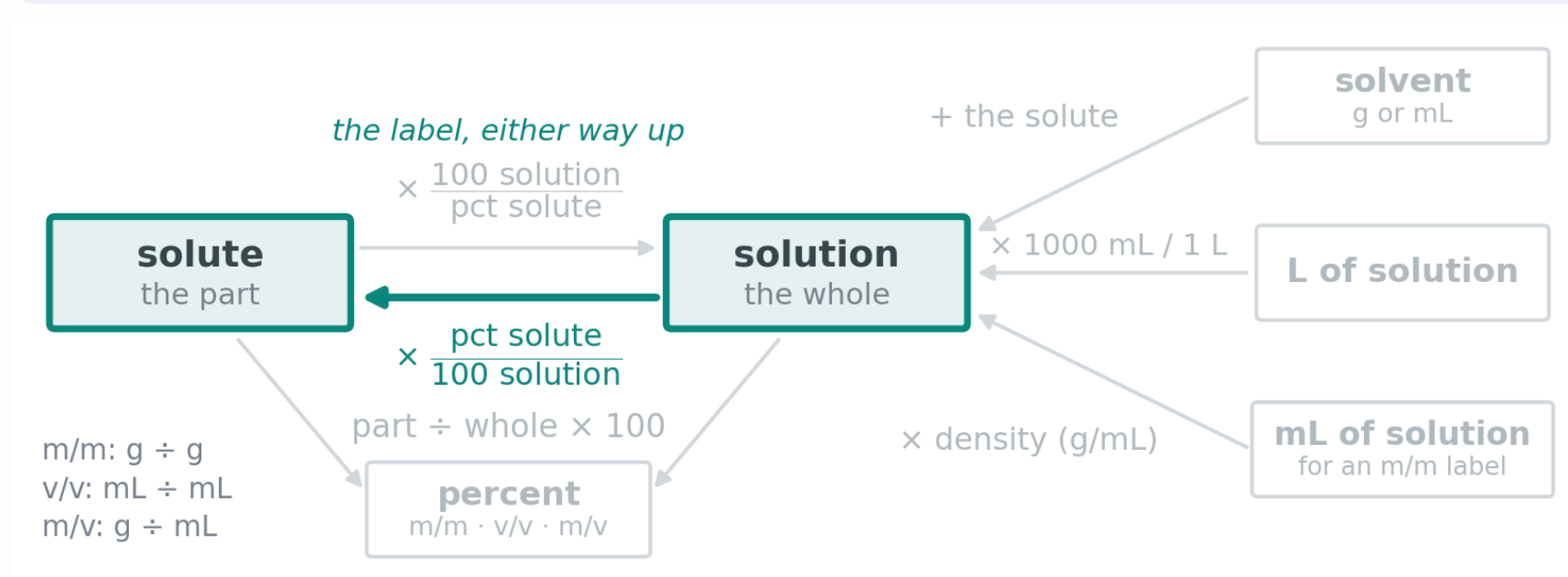
Step 4 · Multiply and check

$$250.0 \text{ g solution} \times \frac{3.0 \text{ g H}_2\text{O}_2}{100 \text{ g solution}} = 7.5 \text{ g H}_2\text{O}_2$$

 **Sense check:** The label promises 3.0 g in every 100 g, and 250.0 g is two and a half hundreds: 7.5 g of H₂O₂. The other 242.5 g is water. ✓

✓ Worked example 2: the route on the map

3.0% (m/m) hydrogen peroxide: $3.0 \text{ g H}_2\text{O}_2 = 100 \text{ g solution}$
 given: 250.0 g solution · found: $7.5 \text{ g H}_2\text{O}_2$



Sense check: The given is already the label's whole, grams of solution. One move: the label factor carries it to the solute. ✓

Your turn: rubbing alcohol

70.0% (v/v) isopropyl alcohol: 70.0 mL alcohol = 100 mL solution
 given: 350.0 mL solution · wanted: mL alcohol

A full **350.0 mL** bottle of rubbing alcohol is **70.0% (v/v)** isopropyl alcohol.

$$350.0 \text{ mL solution} \times \frac{\text{_____ mL alcohol}}{\text{_____ mL solution}} = \text{_____ mL alcohol}$$

Fill the factor from the label so mL solution cancels, then compute.

$$350.0 \text{ mL solution} \times \frac{70.0 \text{ mL alcohol}}{100 \text{ mL solution}} = \mathbf{245 \text{ mL alcohol}}$$

⚠ Where this goes wrong

Dividing by the solvent alone. 25.0 g of sucrose in 100.0 g of water: $25.0 / 100.0 \times 100 = 25.0\%$. The sucrose is part of the whole, so the denominator is $25.0 + 100.0 = 125.0$ g of solution: 20.0%.

Flipping part and whole. $125.0 / 25.0 \times 100 = 500.$, above 100. A percent concentration never tops 100, because the part never outweighs its whole. Part over whole gives 20.0%.

Dropping the $\times 100$. $25.0 / 125.0 = 0.200$, the decimal fraction, not a percent. Per hundred: 20.0%.

Using the raw percent in a chain. For 250.0 g of a 3.0% (m/m) solution, $250.0 \times 3.0 = 750$, one hundred times too much. The label means 3.0 g per 100 g of solution; the factor $3.0 \text{ g} / 100 \text{ g}$ gives 7.5 g.



Practice 2

20.0 g KBr dissolved in 230.0 g water

given: 20.0 g solute · 230.0 g solvent · wanted: m/m %

A stockroom solution is prepared by dissolving **20.0 g of potassium bromide (KBr)** in **230.0 g of water**. What is the mass percent of KBr?

- A. 8.00
- B. 8.70
- C. 1.25×10^3
- D. 0.0800

✓ Practice 2: answer A

20.0 g KBr + 230.0 g water = 250.0 g solution

given: 20.0 g solute · 230.0 g solvent · wanted: m/m %

$$\frac{20.0 \text{ g KBr}}{250.0 \text{ g solution}} \times 100 = 8.00\% \text{ (m/m), answer A}$$

B divided by the water alone: $20.0 / 230.0 \times 100 = 8.70$. C flipped part and whole: $250.0 / 20.0 \times 100 = 1.25 \times 10^3$, and no percent concentration tops 100. D stopped at the decimal fraction: $20.0 / 250.0 = 0.0800$; the $\times 100$ makes it 8.00 per hundred.

 **Sense check:** Every 100 g of solution carries 8.00 g of KBr, and the whole 250.0 g carries two and a half times that: 20.0 g. ✓



Worked example 3: solution volume from a solute mass

Step 1 • Name part and whole

0.90% (m/v) saline: 0.90 g NaCl = 100 mL solution
part: NaCl • whole: solution • given: 4.5 g NaCl • wanted: mL solution

Normal saline is **0.90% (m/v)**: grams of NaCl per 100 milliliters of solution. An IV order calls for **4.5 g of NaCl**. What volume of saline delivers it?

Set it up so g NaCl cancels.

✓ Worked example 3: solution

0.90% (m/v) saline: 0.90 g NaCl = 100 mL solution
 part: NaCl · whole: solution · given: 4.5 g NaCl · wanted: mL solution

Step 2 · Match the units

Step 3 · Write the fraction

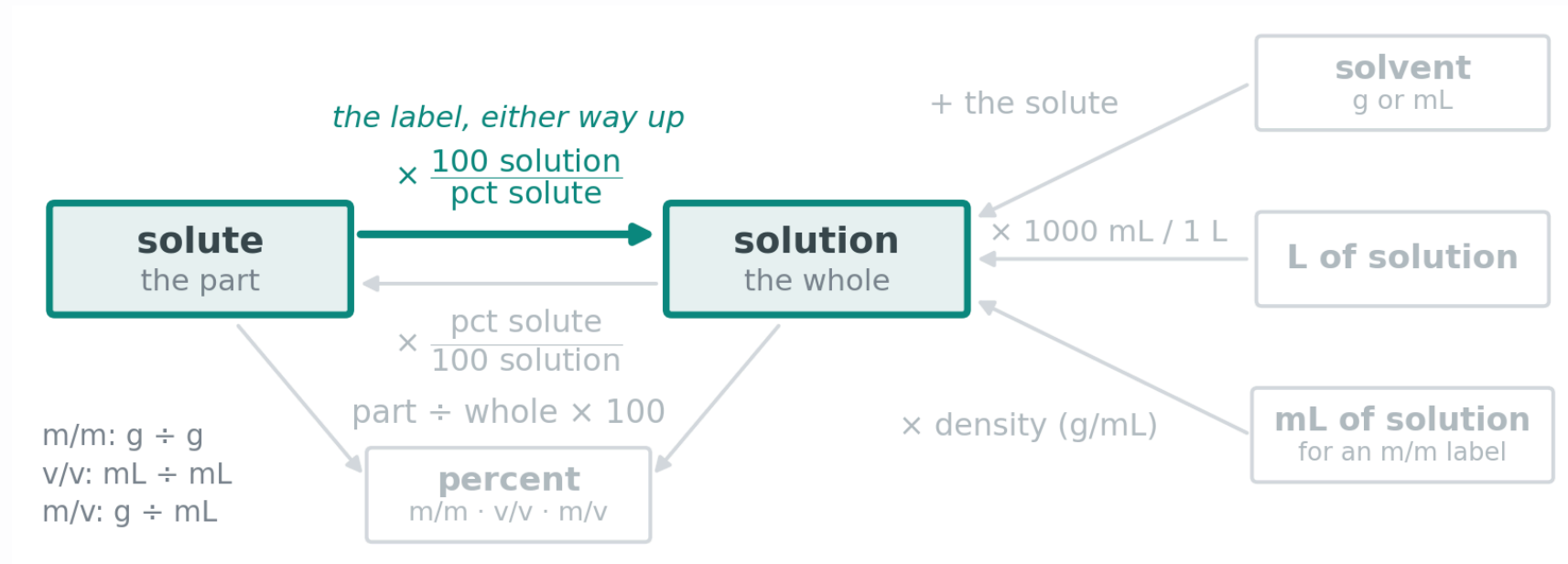
Step 4 · Multiply and check

$$4.5 \text{ g NaCl} \times \frac{100 \text{ mL solution}}{0.90 \text{ g NaCl}} = 5.0 \times 10^2 \text{ mL solution}$$

 **Sense check:** Every 100 mL of saline carries 0.90 g of NaCl, and 4.5 g is five of those hundreds: 5.0×10^2 mL, one standard IV bag. ✓

✓ Worked example 3: the route on the map

0.90% (m/v) saline: 0.90 g NaCl = 100 mL solution
 given: 4.5 g NaCl · found: 5.0×10^2 mL solution



🧠 **Sense check:** The given is the part, so the label factor runs flipped: solute to solution. One move. ✓



Practice 3

IV bag: 15.0% (m/v) mannitol
given: 0.500 L solution · wanted: g mannitol

An IV bag holds **0.500 L** of **15.0% (m/v)** mannitol solution. How many grams of mannitol does it deliver?

- A. 65.2
- B. 0.0750
- C. 3.33×10^3
- D. 75.0
- E. 7.50×10^3

✓ Practice 3: answer D

15.0% (m/v): 15.0 g mannitol = 100 mL solution
 given: 0.500 L solution · wanted: g mannitol · the label's whole is mL

Two conversion factors are needed.

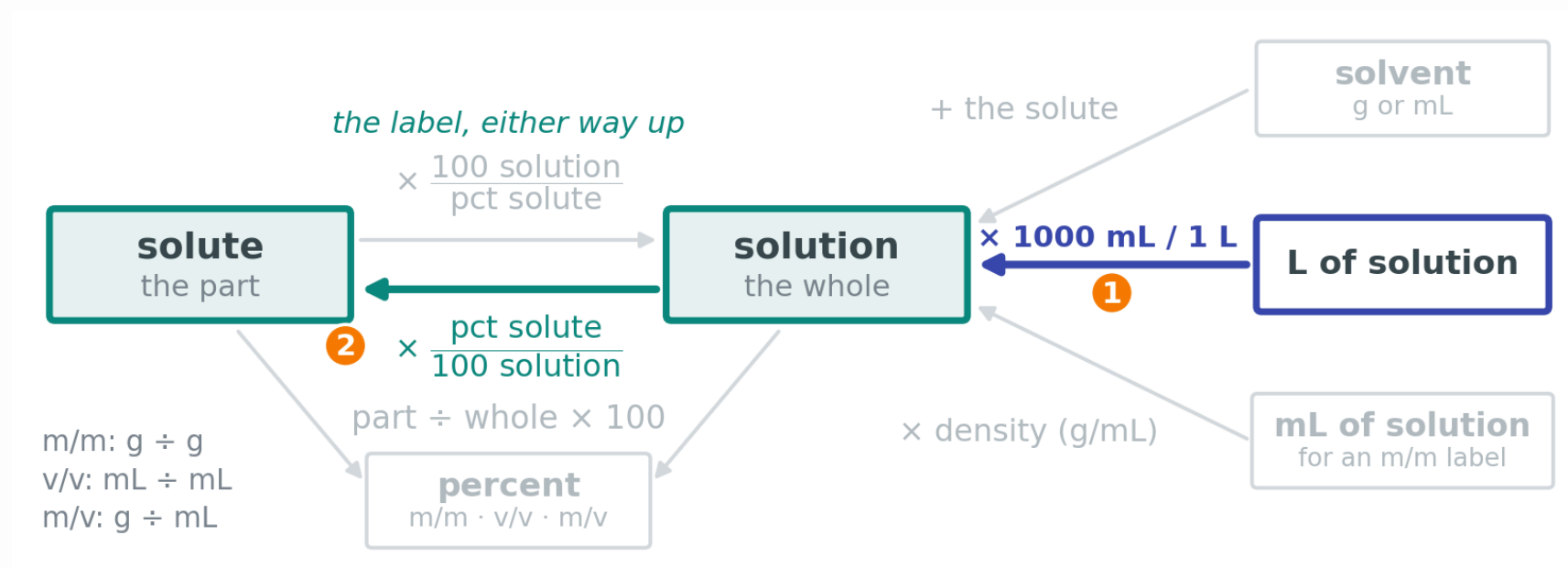
$$0.500 \text{ L soln} \times \frac{1000 \text{ mL soln}}{1 \text{ L soln}} \times \frac{15.0 \text{ g mannitol}}{100 \text{ mL soln}} = \mathbf{75.0 \text{ g mannitol (answer D)}}$$

A read the label as 15.0 g per 100 mL of water, as if 115 parts made the solution: $500. \times 15.0 / 115 = 65.2$. B skipped liters to milliliters: $0.500 \times 15.0 / 100 = 0.0750$. C flipped the label factor: $500. \times 100 / 15.0 = 3.33 \times 10^3$. E used the raw percent: $500. \times 15.0 = 7.50 \times 10^3$.

 **Sense check:** 15.0 g in every 100 mL, and 500. mL is five hundreds: 75.0 g. ✓

✓ Practice 3: the route on the map

15.0% (m/v): 15.0 g mannitol = 100 mL solution
given: 0.500 L solution · found: 75.0 g mannitol



🧠 **Sense check:** The label's whole is milliliters, so liters convert first. Then the label factor carries solution to solute. ✓

 Practice 4

concentrated aqueous ammonia: 28.0% (m/m) NH_3 · density 0.900 g/mL
given: 250. mL solution · wanted: g NH_3

A stockroom bottle holds **250. mL** of concentrated aqueous ammonia, **28.0% (m/m)** NH_3 , with a density of **0.900 g/mL**. How many grams of NH_3 does it hold?

- A. 70.0
- B. 63.0
- C. 225
- D. 77.8

✓ Practice 4: answer B

28.0% (m/m): $28.0 \text{ g NH}_3 = 100 \text{ g solution} \cdot 0.900 \text{ g solution} = 1 \text{ mL solution}$

m/m pairs grams with grams: the density turns mL of solution into g first

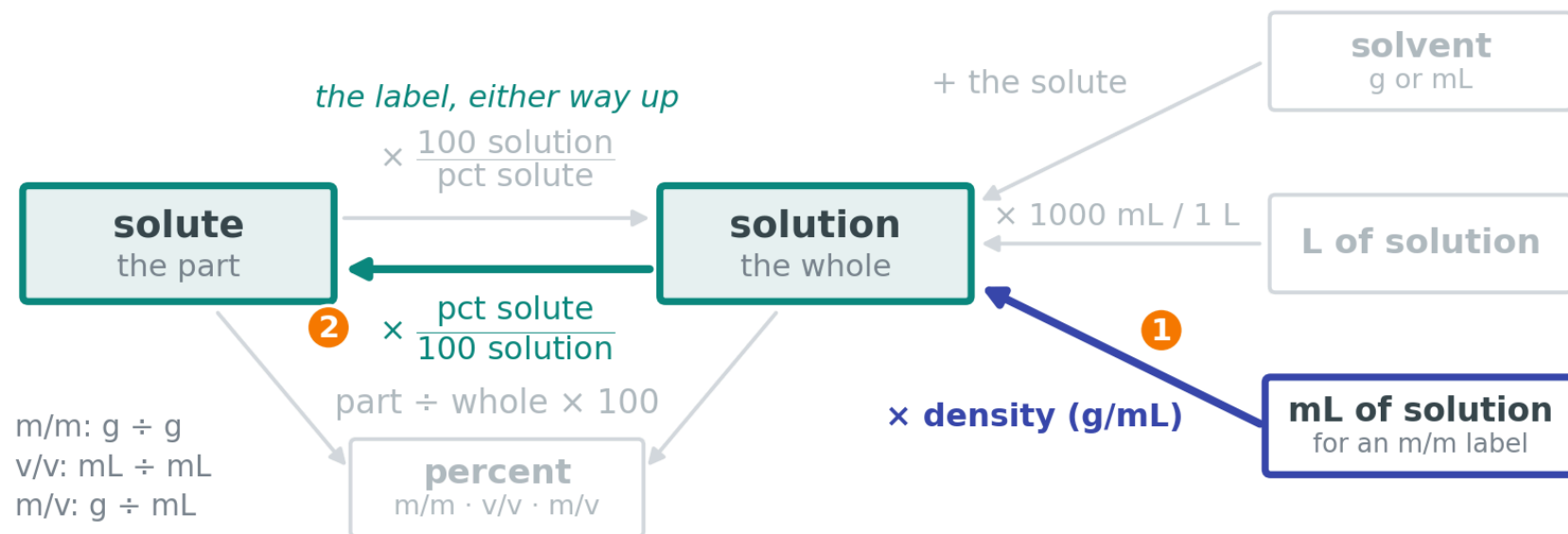
$$250. \cancel{\text{ mL soln}} \times \frac{0.900 \cancel{\text{ g soln}}}{1 \cancel{\text{ mL soln}}} \times \frac{28.0 \text{ g NH}_3}{100 \cancel{\text{ g soln}}} = 63.0 \text{ g NH}_3 \text{ (answer B)}$$

A skipped the density: $250. \times 28.0 / 100 = 70.0$ treats mL of solution as grams. C stopped halfway: $250. \times 0.900 = 225 \text{ g}$ is the whole solution, not the NH_3 in it. D flipped the density: $250. \div 0.900 = 278 \text{ g}$, then $\times 28.0 / 100 = 77.8$.

 **Sense check:** The liquid is lighter than water, so 250. mL weighs less than 250 g, and the NH_3 lands below the 70.0 g a density of 1 would give: 63.0 g. ✓

✓ Practice 4: the route on the map

28.0% (m/m) NH_3 · density 0.900 g/mL
given: 250. mL solution · found: 63.0 g NH_3



Sense check: An m/m label counts grams of solution, so the density turns milliliters into grams first. Then the label factor. ✓



Check yourself

1. A hospital dextrose bag is labeled **5.0% (m/v)**. State the equality the label declares, then write the factor that converts milliliters of solution to grams of dextrose.
2. **10.0 g of NaOH** dissolves in **90.0 g of water**. What mass belongs in the denominator of the mass percent, and what is its value?

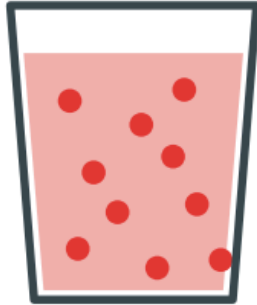
Water-quality reports push the same fraction further: parts per million and parts per billion, for solutes far too dilute to reach one part in a hundred.

5 • Molarity

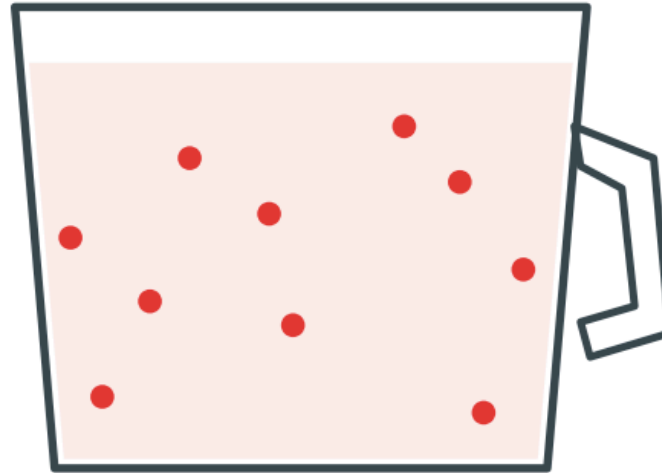
Calculate the molarity of a solution from the amount of solute and the volume of solution, and use it as a conversion factor between solution volume and moles of solute, including the grams to weigh out to prepare a solution.

Same mix, different strength

one spoonful of mix in each; only the water differs



a small glass
tastes strong

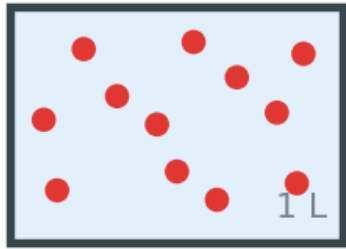


a full pitcher
barely tastes at all

One spoonful of drink mix in a small glass tastes strong. The same spoonful in a full pitcher barely tastes at all. Only the water changed.

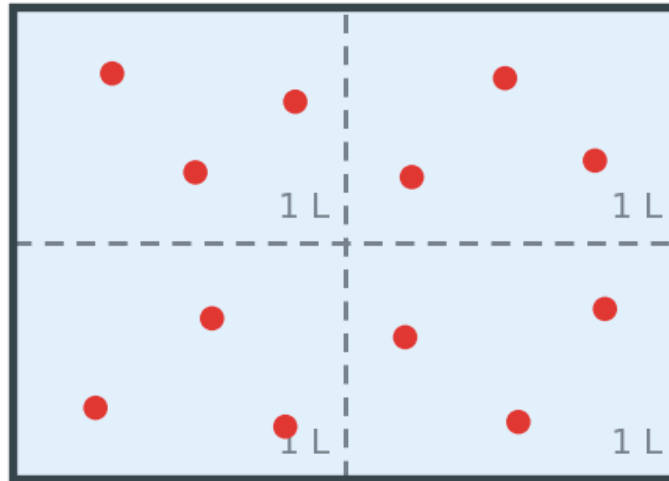
Amount per liter, not total amount

same total amount, a different amount in each liter



12 particles in 1 L

12 per liter



the same 12 particles in 4 L

3 per liter

Spread the same twelve particles through four liters instead of one, and each liter holds three. Taste, color, dose, and reactivity follow the amount in each liter, not the total.

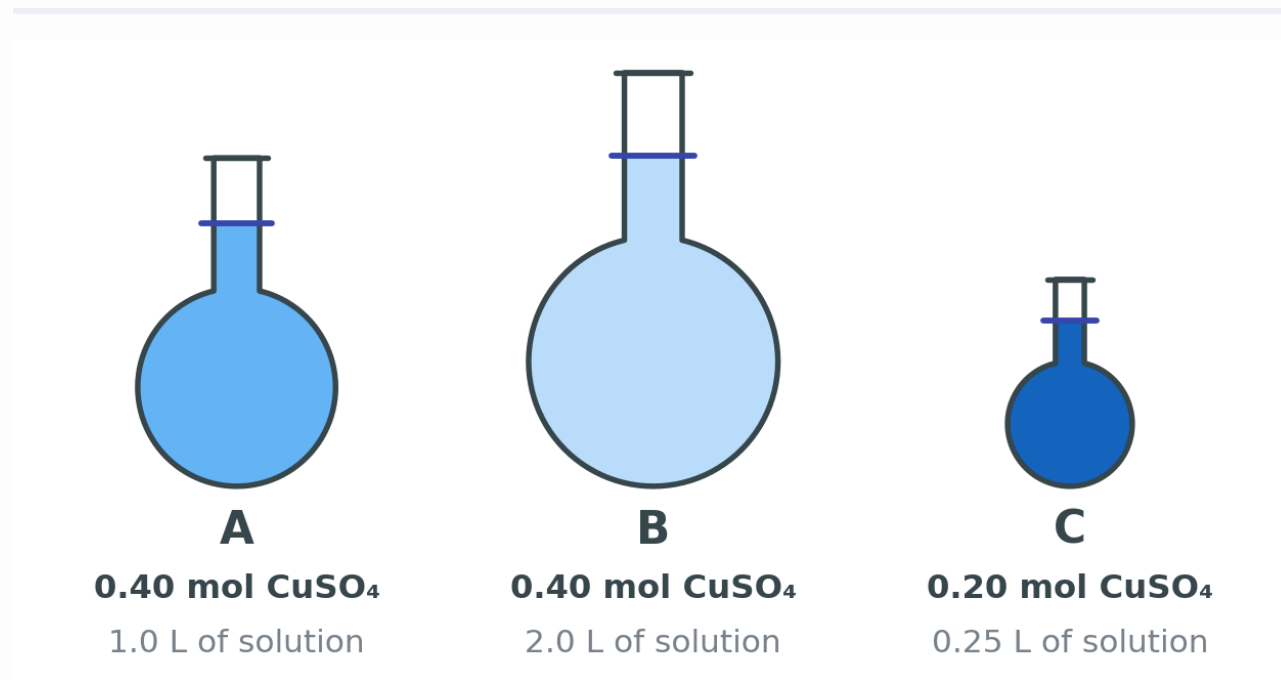
Solute, solvent, solution

solute + solvent → solution

dissolved substance · dissolving medium · uniform mixture

Sugar stirred into water spreads evenly through it. The sugar is the **solute**, the water the **solvent**, and the mixture a **solution**. Concentration states how much solute each volume of solution carries.

Comparing three solutions



Concentration compares the amount of solute to the amount of solution. Rank the three copper(II) sulfate solutions: which is most concentrated, and which is most dilute?

mol $\div$ L: A $0.40 \div 1.0 = 0.40$ · B $0.40 \div 2.0 = 0.20$ · C $0.20 \div 0.25 = 0.80$
most concentrated: C · most dilute: B · C holds the least CuSO_4 in total but the most in each liter

Molarity: moles per liter of solution

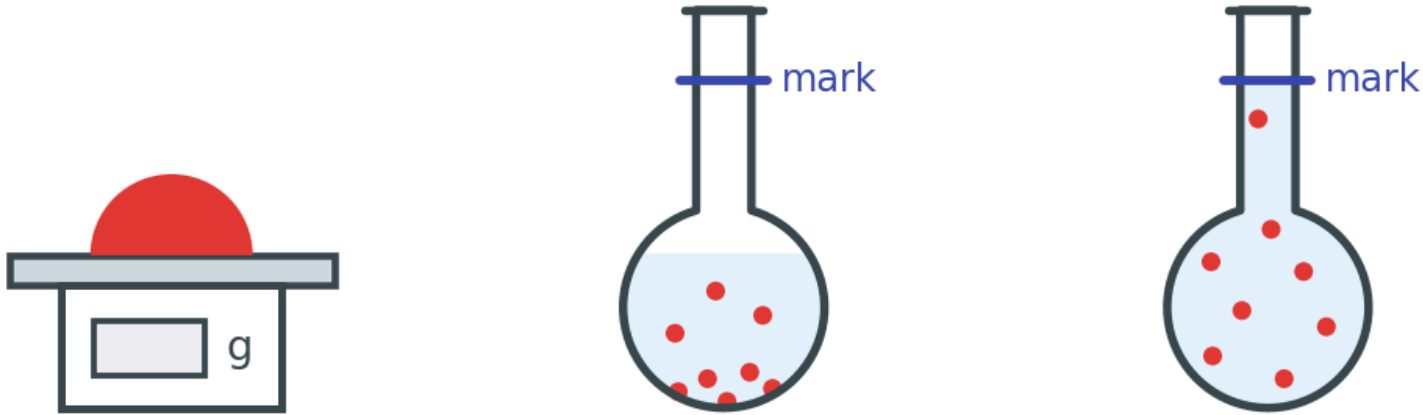
$$M = \text{mol solute} \div \text{L solution}$$

6.0 M HCl: every liter of the solution carries 6.0 mol HCl · read "six molar"

Molarity, symbol M, counts the moles of solute in each liter of solution: a rate, not a total. Reactions consume particles, not grams; two solutions at the same g/L can carry different particle counts.

Liters of solution, not water added

the mark reads the volume of solution, solute included



1 · weigh the solute 2 · add water, dissolve 3 · fill to the mark

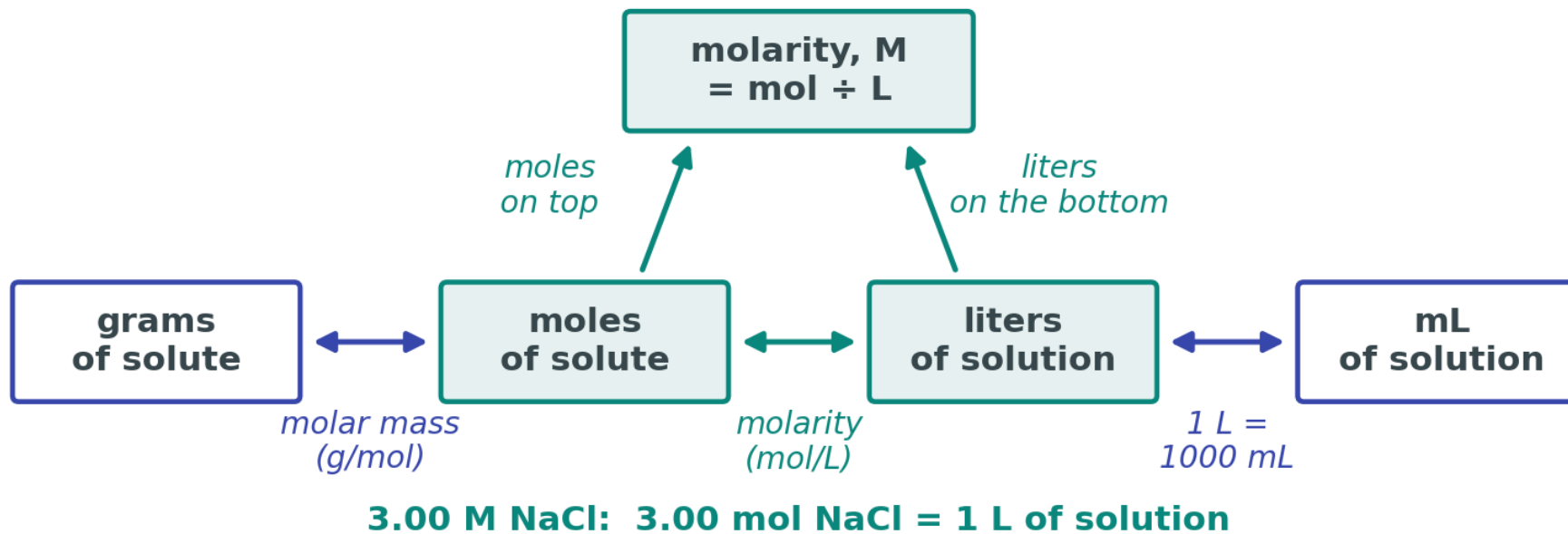
The solute takes up room. Prepare the solution in a volumetric flask: solute in first, then water to the mark. The mark reads the volume of finished solution, solute included.



The method

1. **Grams $\rightarrow$ moles:** convert the solute's mass with its molar mass.
2. **mL $\rightarrow$ L:** molarity counts liters of solution.
3. **Divide moles by liters:** the quotient is the molarity, in mol/L.

The molarity map



Divide moles by liters to find a molarity. Once known, the molarity converts between moles of solute and liters of solution. The molar mass reaches grams.



Guided example: a sports drink

$$M = \text{mol solute} \div \text{L solution}$$

given: 0.300 mol glucose · 1.50 L of solution · wanted: M

A sports drink is mixed so that **1.50 L** of it holds **0.300 mol of glucose**. What is its molarity?

Check each given against the three method steps, then divide.

✓ Guided example: solution

$$M = \text{mol solute} \div \text{L solution}$$

given: 0.300 mol glucose · 1.50 L of solution · wanted: M

Step 1 · Grams → moles

Step 2 · mL → L

Step 3 · Divide moles by liters

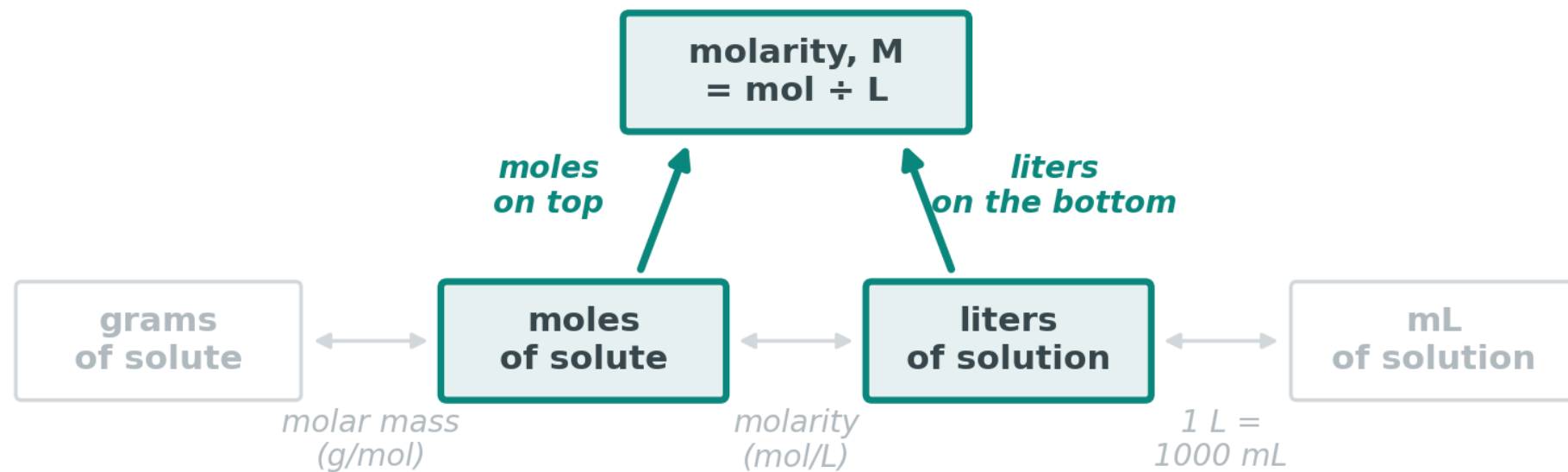
$$M = \frac{0.300 \text{ mol glucose}}{1.50 \text{ L soln}} = 0.200 \text{ M glucose}$$

🧠 **Sense check:** 1.50 L holds 0.300 mol, so each half liter holds 0.100 mol and a full liter holds 0.200 mol. 0.200 M ✓

✓ Guided example: the route on the map

$$M = \text{mol solute} \div \text{L solution}$$

given: 0.300 mol glucose · 1.50 L of solution · found: 0.200 M glucose



🧠 **Sense check:** Moles and liters were both in hand, so the route is one move: moles on top, liters on the bottom. ✓



Practice 1

$$M = \text{mol solute} \div \text{L solution}$$

measured: 0.270 mol sucrose · 1.15 L of water · 1.20 L after dissolving

A student stirs **0.270 mol** of sucrose into **1.15 L** of water. Once it dissolves, the solution measures **1.20 L**. What is the molarity of the sucrose solution?

- A. 0.235
- B. 4.44
- C. 0.324
- D. 0.225

✓ Practice 1: answer D

$$M = \text{mol solute} \div \text{L solution}$$

0.270 mol sucrose · 1.20 L of solution · the 1.15 L of water is not the solution volume

$$M = \frac{0.270 \text{ mol sucrose}}{1.20 \text{ L soln}} = \mathbf{0.225 \text{ M sucrose (answer D)}}$$

A divided by the water added: $0.270 \div 1.15 = 0.235$, but the dissolved sugar takes up room too. B flipped the fraction: $1.20 \div 0.270 = 4.44$, in L/mol. C multiplied: $0.270 \times 1.20 = 0.324$, in mol·L.

🧠 **Sense check:** A bit more than a liter holds 0.270 mol, so one liter holds a bit less: $0.225 < 0.270$ ✓



Worked example 1: molarity from moles and volume

$$M = \text{mol solute} \div \text{L solution}$$

given: 0.350 mol NaCl · 500.0 mL of solution · wanted: M

A stockroom bottle is prepared with **0.350 mol of NaCl** dissolved in enough water to make **500.0 mL of solution**. Find the molarity for the label.

Set it up: get the volume into liters, then divide.

✓ Worked example 1: solution

$$M = \text{mol solute} \div \text{L solution}$$

given: 0.350 mol NaCl · 500.0 mL of solution · wanted: M

Step 1 · Grams → moles

Step 2 · mL → L

$$500.0 \text{ mL} = 0.5000 \text{ L}$$

1000 mL = 1 L · the definition counts liters of solution

Step 3 · Divide moles by liters

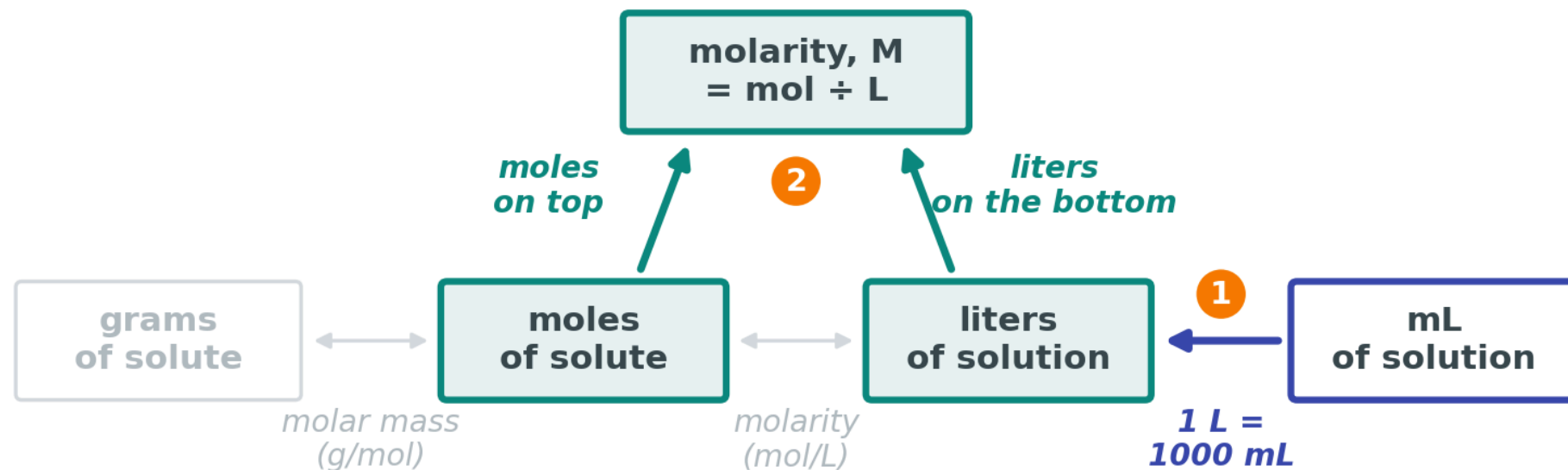
$$M = \frac{0.350 \text{ mol NaCl}}{0.5000 \text{ L soln}} = 0.700 \text{ M NaCl}$$

🧠 **Sense check:** Half a liter carries 0.350 mol, so a full liter carries twice that: 0.700 mol. 0.700 M ✓

✓ Worked example 1: the route on the map

$$M = \text{mol solute} \div \text{L solution}$$

given: 0.350 mol NaCl · 500.0 mL of solution · found: 0.700 M NaCl



🧠 **Sense check:** Two moves: 500.0 mL becomes 0.5000 L, then moles on top, liters on the bottom. ✓



Worked example 2: molarity from grams

$$M = \text{mol solute} \div \text{L solution}$$

given: 9.35 g KCl · water to the 250.0 mL mark · molar mass KCl 74.55 g/mol

A student measures out **9.35 g of KCl** (74.55 g/mol), transfers it to a volumetric flask, and adds water to the **250.0 mL** mark. Find the molarity.

A common first attempt: divide the moles by 250. Test the units.

✓ Worked example 2: solution

$$M = \text{mol solute} \div \text{L solution}$$

given: 9.35 g KCl (74.55 g/mol) · water to the 250.0 mL mark · wanted: M

Step 1 · Grams → moles

A common first attempt

$$M = \frac{0.1254 \text{ mol KCl}}{250 \text{ mL soln}} = 0.000502 \text{ mol/mL} \times$$

Step 2 · mL → L

Step 3 · Divide moles by liters

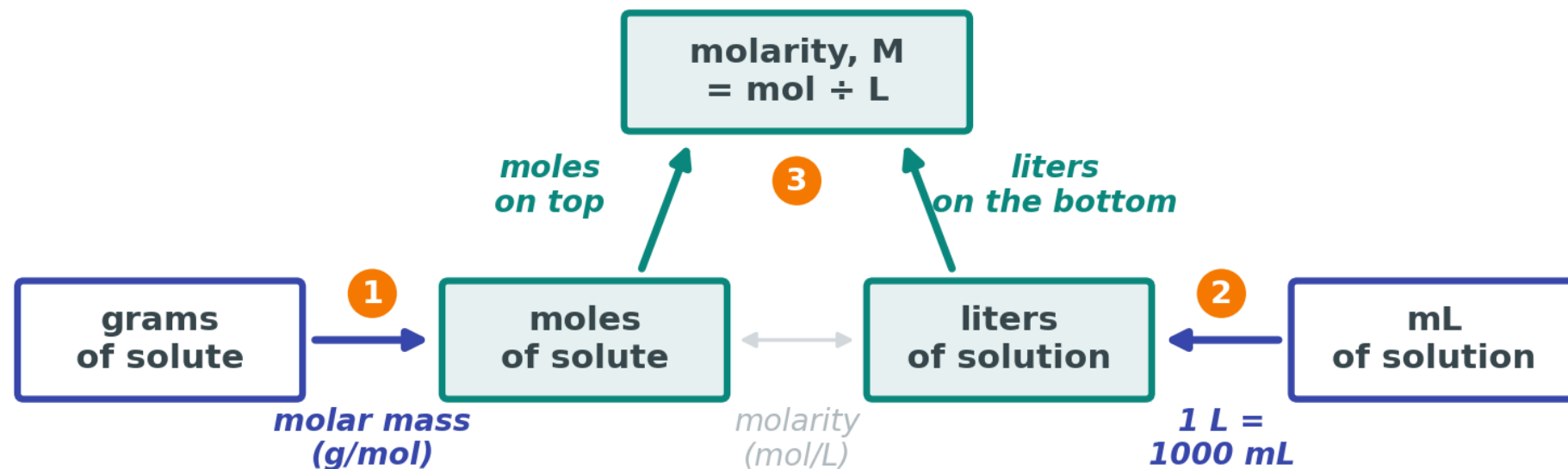
$$9.35 \text{ g KCl} \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} = 0.1254 \text{ mol, then } \frac{0.1254 \text{ mol KCl}}{0.2500 \text{ L soln}} = \mathbf{0.502 \text{ M KCl}}$$

🧠 **Sense check:** A quarter liter carries 0.1254 mol, so a full liter carries four times that: 0.502 mol. 0.502 M ✓

✓ Worked example 2: the route on the map

$$M = \text{mol solute} \div \text{L solution}$$

given: 9.35 g KCl (74.55 g/mol) · 250.0 mL of solution · found: 0.502 M KCl



🧠 **Sense check:** Three moves: the molar mass turns grams into moles, 250.0 mL becomes 0.2500 L, then divide. ✓

Take-home: convert mL to L first

$$M = \frac{0.1254 \text{ mol}}{250 \text{ mL}} = 0.000502 \text{ mol/mL} \times (\text{not a molarity})$$

$$M = \frac{0.1254 \text{ mol}}{0.2500 \text{ L}} = \mathbf{0.502 \text{ M} \checkmark}$$

Molarity is defined per liter of solution. A volume left in milliliters gives a result 1000 times too small. Convert every volume to liters before it enters the definition.

Your turn: sodium carbonate

$$M = \text{mol solute} \div \text{L solution}$$

given: 21.2 g Na₂CO₃ (105.99 g/mol) · water to the 500.0 mL mark · wanted: M

A wash solution is prepared: **21.2 g of Na₂CO₃** (105.99 g/mol) in a volumetric flask, water to the **500.0 mL** mark.

$$21.2 \text{ g Na}_2\text{CO}_3 \times \frac{1 \text{ mol Na}_2\text{CO}_3}{\text{ } \text{g Na}_2\text{CO}_3} = \text{ } \text{mol, then } \frac{\text{ } \text{mol Na}_2\text{CO}_3}{\text{ } \text{L soln}} = \text{ } \text{M}$$

Fill the molar mass, the moles, and the liters, then compute the molarity.

$$21.2 \text{ g Na}_2\text{CO}_3 \times \frac{1 \text{ mol Na}_2\text{CO}_3}{105.99 \text{ g Na}_2\text{CO}_3} = 0.200 \text{ mol, then } \frac{0.200 \text{ mol Na}_2\text{CO}_3}{0.5000 \text{ L soln}} = \mathbf{0.400 \text{ M Na}_2\text{CO}_3}$$

⚠ Where this goes wrong

9.35 g KCl (74.55 g/mol) · water to the 250.0 mL mark

correct: $9.35 \text{ g} \rightarrow 0.1254 \text{ mol} \cdot 250.0 \text{ mL} \rightarrow 0.2500 \text{ L} \cdot \text{M} = 0.502 \text{ M}$

Dividing by milliliters. $0.1254 \text{ mol} \div 250 \text{ mL} = 0.000502$, and the unit is mol/mL, 1000 times too small. Molarity divides by liters of solution: $0.1254 \div 0.2500 = 0.502 \text{ M}$.

Dividing grams by liters. $9.35 \div 0.2500 = 37.4$ is a mass concentration in g/L, not a molarity. Molarity counts moles of solute: $9.35 \text{ g} \div 74.55 \text{ g/mol} = 0.1254 \text{ mol}$ first.

Multiplying by the molar mass. $9.35 \times 74.55 = 697$ carries units of g^2/mol , which is not a mole count. Grams $\rightarrow$ moles divides by the molar mass: $9.35 \div 74.55 = 0.1254 \text{ mol}$.



Practice 2

$$M = \text{mol solute} \div \text{L solution}$$

molar mass KNO_3 101.11 g/mol

What is the molarity of **225 mL** of a potassium nitrate solution that contains **34.8 g** of KNO_3 ?

- A. 0.344
- B. 1.53
- C. 0.00153
- D. 15.5
- E. 155

✓ Practice 2: answer B

$$M = \text{mol solute} \div L \text{ solution}$$

$$34.8 \text{ g KNO}_3 (101.11 \text{ g/mol}) \cdot 225 \text{ mL of solution} = 0.225 \text{ L} \cdot \text{wanted: M}$$

$$34.8 \text{ g KNO}_3 \times \frac{1 \text{ mol KNO}_3}{101.11 \text{ g KNO}_3} = 0.344 \text{ mol, then } \frac{0.344 \text{ mol KNO}_3}{0.225 \text{ L soln}}$$

= 1.53 M KNO₃ (answer B)

A stopped at moles: $34.8 \div 101.11 = 0.344 \text{ mol}$ is the amount in the sample. C divided by milliliters: $0.344 \div 225 = 0.00153$, in mol/mL. D found the percent: $34.8 \div 225 \times 100 = 15.5\%$ (m/v). E divided grams by liters: $34.8 \div 0.225 = 155 \text{ g/L}$.

 **Sense check:** A bit under a quarter liter carries 0.344 mol, so a full liter carries a bit over four times that, over 1.38 mol. 1.53 M ✓



Practice 3

$$M = \text{mol solute} \div \text{L solution}$$

measured: 2.000 L batch · 25.00 mL portion dried · 2.38 g KBr residue · molar mass KBr 119.00 g/mol

A technician prepares a **2.000 L** batch of KBr solution. A **25.00 mL** portion of it is evaporated to dryness, and the residue weighs **2.38 g**. What is the molarity of the batch?

- A. 0.800
- B. 0.0100
- C. 0.0200
- D. 0.000800
- E. 0.000500

✓ Practice 3: answer A

$$M = \text{mol solute} \div L \text{ solution}$$

25.00 mL portion = 0.02500 L · 2.38 g KBr (119.00 g/mol) in it · the portion has the batch's molarity

$$2.38 \text{ g KBr} \times \frac{1 \text{ mol KBr}}{119.00 \text{ g KBr}} = 0.0200 \text{ mol, then } \frac{0.0200 \text{ mol KBr}}{0.02500 \text{ L soln}}$$

= **0.800 M KBr (answer A)**

B divided by the whole batch: $0.0200 \div 2.000 = 0.0100$, but the 2.38 g came from only 25.00 mL. C stopped at moles: 0.0200 mol is the amount in the portion. D divided by milliliters: $0.0200 \div 25.00 = 0.000800$, in mol/mL. E multiplied by the volume instead of dividing: $0.0200 \times 0.02500 = 0.000500$.

 **Sense check:** A uniform solution has the same molarity in every portion. A fortieth of a liter carries 0.0200 mol, so a full liter carries forty times that: 0.800 mol. 0.800 M ✓

A molarity is an equality

0.450 mol KOH = 1 L of solution
the label "0.450 M KOH" states this equality

Every equality gives a conversion factor. This one converts between volume of solution and moles of solute:

$$\frac{0.450 \text{ mol KOH}}{1 \text{ L soln}} \text{ or } \frac{1 \text{ L soln}}{0.450 \text{ mol KOH}}$$

Write it so the **given unit cancels**.



Worked example 3: liters from grams

3.00 mol NaCl = 1 L of solution

given: 351 g NaCl (58.44 g/mol) · 3.00 M NaCl · wanted: L of solution · route: g → mol → L

How many liters of **3.00 M NaCl** solution can be made from **351 g of NaCl**?

Count the factors on the route: g → mol → L.

✓ Worked example 3: solution

3.00 mol NaCl = 1 L of solution

given: 351 g NaCl (58.44 g/mol) · wanted: L of solution · route: g → mol → L

Step 1 · Grams → moles

$$351 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 6.01 \text{ mol NaCl}$$

Write the molarity fraction

Multiply and check

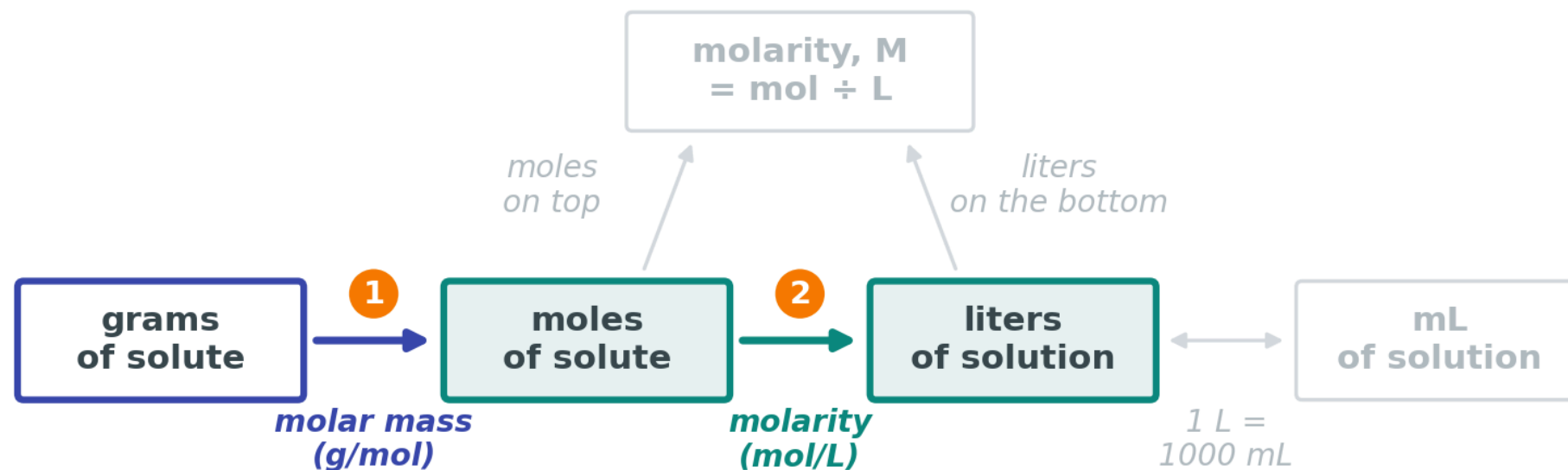
$$351 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \times \frac{1 \text{ L soln}}{3.00 \text{ mol NaCl}} = 2.00 \text{ L soln}$$

🧠 **Sense check:** Each liter carries 3.00 mol, and 351 g is 6.01 mol: two liters' worth. 2.00 L ✓

✓ Worked example 3: the route on the map

3.00 mol NaCl = 1 L of solution

given: 351 g NaCl · 3.00 M NaCl · found: 2.00 L of solution



🧠 **Sense check:** Two moves: the molar mass reaches moles, then the molarity, liters over moles, reaches liters. ✓



Worked example 4: volume from mass of solute

0.450 mol KOH = 1 L of solution

given: 15.1 g KOH (56.11 g/mol) · wanted: mL of solution · route: g → mol → L → mL

A procedure calls for **15.1 g of KOH**, and the shelf stocks **0.450 M KOH solution**. What volume of that solution, in milliliters, delivers the 15.1 g?

Count the factors on the route: g → mol → L → mL.

✓ Worked example 4: solution

0.450 mol KOH = 1 L of solution

given: 15.1 g KOH (56.11 g/mol) · wanted: mL of solution · route: g → mol → L → mL

Step 1 · Grams → moles

$$15.1 \text{ g KOH} \times \frac{1 \text{ mol KOH}}{56.11 \text{ g KOH}} = 0.269 \text{ mol KOH}$$

Multiply and check

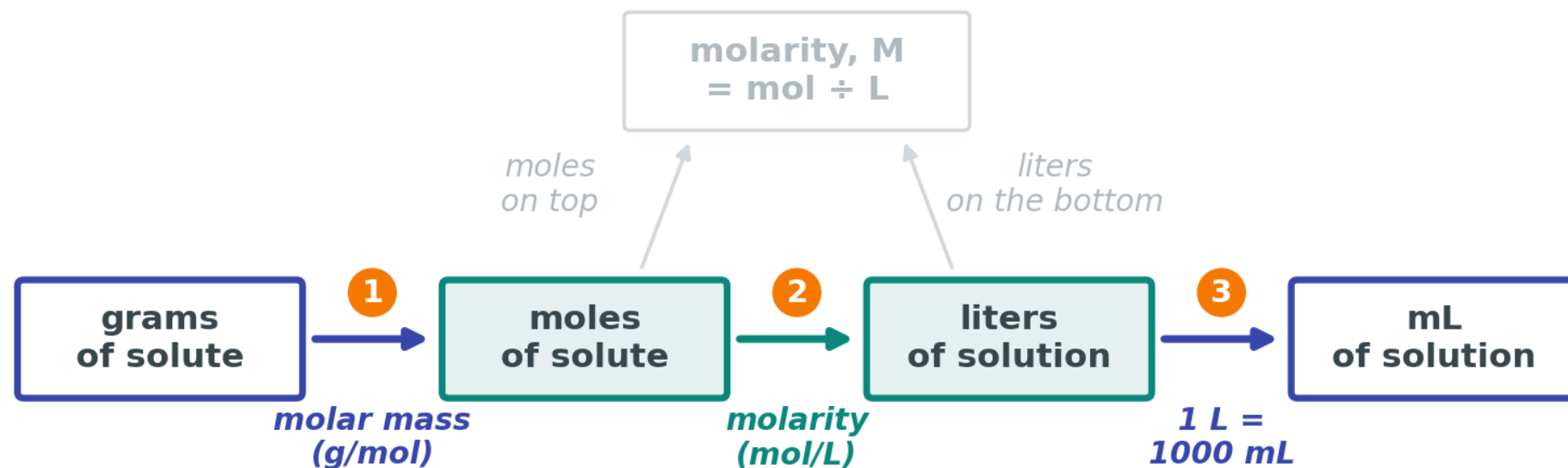
$$15.1 \text{ g KOH} \times \frac{1 \text{ mol KOH}}{56.11 \text{ g KOH}} \times \frac{1 \text{ L soln}}{0.450 \text{ mol KOH}} \times \frac{1000 \text{ mL soln}}{1 \text{ L soln}} = 598 \text{ mL soln}$$

 **Sense check:** Each liter carries 0.450 mol, and 0.269 mol is wanted, a bit over half a liter: 598 mL ✓

✓ Worked example 4: the route on the map

0.450 mol KOH = 1 L of solution

given: 15.1 g KOH · 0.450 M KOH · found: 598 mL of solution



🧠 **Sense check:** One move past liters: 1 L = 1000 mL turns 0.598 L into 598 mL. ✓



Practice 4

$0.400 \text{ mol KNO}_3 = 1 \text{ L of solution}$

the label " 0.400 M KNO_3 " states this equality · molar mass KNO_3 101.11 g/mol

A greenhouse feed calls for **10.1 g of KNO_3** , supplied as **0.400 M KNO_3 solution**. What volume, in milliliters, carries that mass?

- A. 0.250
- B. 40.0
- C. 0.0999
- D. 250.

✓ Practice 4: answer D

0.400 mol KNO₃ = 1 L of solution

given: 10.1 g KNO₃ (101.11 g/mol) · wanted: mL of solution · route: g → mol → L → mL

$$10.1 \text{ g KNO}_3 \times \frac{1 \text{ mol KNO}_3}{101.11 \text{ g KNO}_3} \times \frac{1 \text{ L soln}}{0.400 \text{ mol KNO}_3} \times \frac{1000 \text{ mL soln}}{1 \text{ L soln}}$$

= 250. mL (answer D)

A stopped at liters and relabeled: $10.1 \div 101.11 \div 0.400 = 0.250 \text{ L}$, which is 250. mL, not 0.250 mL. B flipped the molarity fraction: $10.1 \div 101.11 \times 0.400 \times 1000 = 40.0$, and the units, mol²/L, are not a volume. C stopped at moles: $10.1 \div 101.11 = 0.0999 \text{ mol}$ is the KNO₃ needed, not the volume that carries it.

🧠 **Sense check:** Each liter carries 0.400 mol, and the target is 0.0999 mol, a quarter of a liter: 250. mL ✓



Worked example 5: preparing a solution from the solid

0.10 mol NaOH = 1 L of solution

given: 1.0 L of 0.10 M NaOH to prepare · NaOH 40.00 g/mol · wanted: g NaOH, then the steps

Describe how to prepare **1.0 L of 0.10 M NaOH** from solid NaOH.

Find the mass first. Count the factors on the route: L → mol → g.

✓ Worked example 5: solution

0.10 mol NaOH = 1 L of solution

given: 1.0 L of 0.10 M NaOH (40.00 g/mol) · wanted: g NaOH · route: L → mol → g

Liters → moles: the molarity

$$1.0 \text{ L soln} \times \frac{0.10 \text{ mol NaOH}}{1 \text{ L soln}} = 0.10 \text{ mol NaOH}$$

Moles → grams: the molar mass

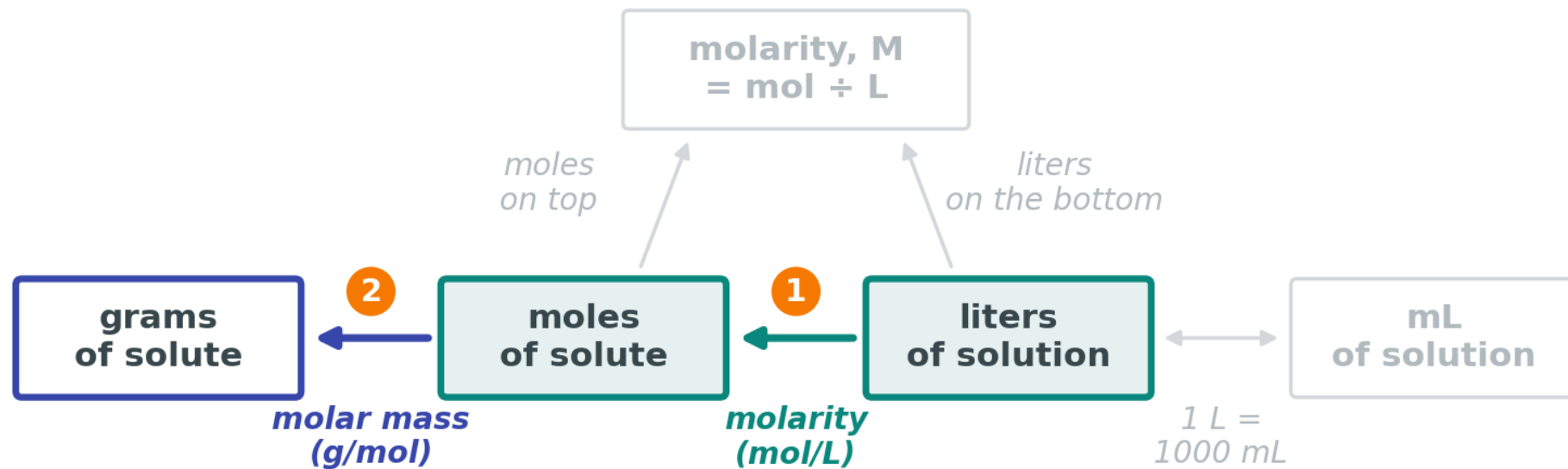
$$1.0 \text{ L soln} \times \frac{0.10 \text{ mol NaOH}}{1 \text{ L soln}} \times \frac{40.00 \text{ g NaOH}}{1 \text{ mol NaOH}} = 4.0 \text{ g NaOH}$$

 **Sense check:** A liter of 0.10 M holds 0.10 mol, a tenth of a mole: a tenth of 40.00 g is 4.0 g ✓

✓ Worked example 5: the route on the map

0.10 mol NaOH = 1 L of solution

given: 1.0 L of 0.10 M NaOH · found: 4.0 g NaOH



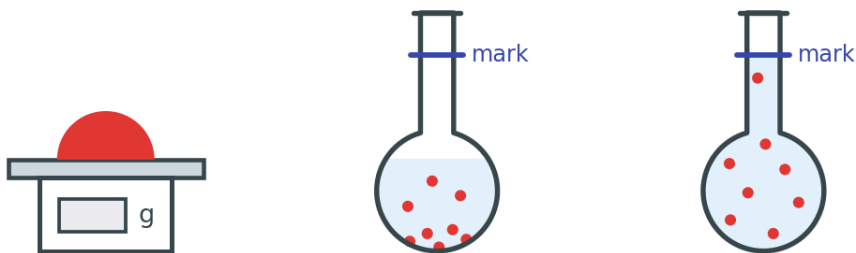
🧠 **Sense check:** Liters in, grams out: the molarity reaches moles, then the molar mass reaches grams. ✓

✓ Worked example 5: making the solution

$0.10 \text{ mol NaOH} = 1 \text{ L of solution}$

found: 4.0 g NaOH · flask: 1.0 L volumetric

the mark reads the volume of solution, solute included



1 · weigh the solute 2 · add water, dissolve 3 · fill to the mark

Weigh 4.0 g of NaOH. Dissolve it in part of the water in a 1.0 L volumetric flask. Add water to the mark, then mix.

🧠 **Sense check:** The mark measures 1.0 L of solution, not 1.0 L of water added. ✓

 Practice 5

0.150 mol Na_2SO_4 = 1 L of solution

molar mass Na_2SO_4 142.04 g/mol

A dye bath calls for **500.0 mL of 0.150 M Na_2SO_4** . What mass of Na_2SO_4 , in grams, must be dissolved to make it?

- A. 21.3
- B. 10.7
- C. 473
- D. 1.07×10^4

✓ Practice 5: answer B

0.150 mol Na₂SO₄ = 1 L of solution

given: 500.0 mL of 0.150 M Na₂SO₄ (142.04 g/mol) · wanted: g Na₂SO₄ · route: mL → L → mol → g

$$500.0 \text{ mL soln} \times \frac{1 \text{ L soln}}{1000 \text{ mL soln}} \times \frac{0.150 \text{ mol Na}_2\text{SO}_4}{1 \text{ L soln}} \times \frac{142.04 \text{ g Na}_2\text{SO}_4}{1 \text{ mol Na}_2\text{SO}_4}$$

= 10.7 g (answer B)

A left out the volume: $0.150 \times 142.04 = 21.3 \text{ g}$ is the mass in a full liter. C flipped the molarity: $0.5000 \div 0.150 \times 142.04 = 473$, and the units do not cancel. D skipped mL → L: $500.0 \times 0.150 \times 142.04 = 1.07 \times 10^4$.

 **Sense check:** Half a liter of 0.150 M holds 0.0750 mol, at about 142 g per mole: about 11 g. 10.7 g ✓



Worked example 6: ethanol in blood

0.080 mol C₂H₆O = 1 L of solution

given: 5.6 L of blood at 0.080 M ethanol · C₂H₆O 46.07 g/mol · wanted: g ethanol, then % (m/v)

A blood level of **0.080 M ethanol** (C₂H₆O) can induce a coma. What total mass of ethanol is in an adult with **5.6 L** of blood at this level? What is the blood alcohol level in % (m/v)?

Find the mass first. Count the factors on the route: L → mol → g.

✓ Worked example 6: solution

0.080 mol C₂H₆O = 1 L of solution

given: 5.6 L of blood at 0.080 M (46.07 g/mol) · wanted: g C₂H₆O, then % (m/v)

Liters → moles: the molarity

$$5.6 \text{ L soln} \times \frac{0.080 \text{ mol C}_2\text{H}_6\text{O}}{1 \text{ L soln}} = 0.45 \text{ mol C}_2\text{H}_6\text{O}$$

Moles → grams: the molar mass

$$5.6 \text{ L soln} \times \frac{0.080 \text{ mol C}_2\text{H}_6\text{O}}{1 \text{ L soln}} \times \frac{46.07 \text{ g C}_2\text{H}_6\text{O}}{1 \text{ mol C}_2\text{H}_6\text{O}} = 21 \text{ g C}_2\text{H}_6\text{O}$$

 **Sense check:** Each liter carries 0.080 mol, about 3.7 g, and 5.6 × 3.7 is about 21. 21 g ✓

✓ Worked example 6: the percent

0.080 mol $\text{C}_2\text{H}_6\text{O}$ = 1 L of solution

found: 21 g $\text{C}_2\text{H}_6\text{O}$ (20.6 unrounded) in 5.6 L of blood · wanted: % (m/v)

Grams per 100 mL: the % (m/v)

Keep the unrounded 20.6 g in the calculator. The blood is the solution: 5.6 L = 5.6×10^3 mL.

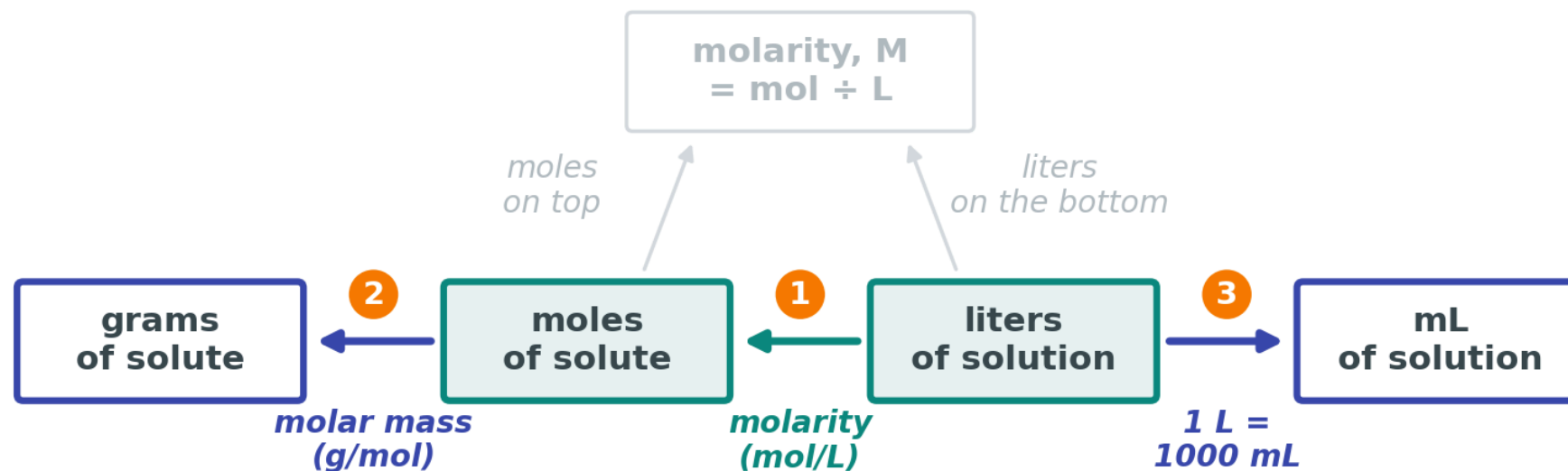
$$\frac{20.6 \text{ g } \text{C}_2\text{H}_6\text{O}}{5.6 \times 10^3 \text{ mL soln}} \times 100 = \mathbf{0.37\% \text{ (m/v)}}$$

 **Sense check:** Each liter carries about 3.7 g, so each 100 mL carries about 0.37 g: 0.37% (m/v) ✓

✓ Worked example 6: the route on the map

0.080 mol C₂H₆O = 1 L of solution

given: 5.6 L of blood at 0.080 M · found: 21 g C₂H₆O · 0.37% (m/v)



🧠 **Sense check:** Molarity, then molar mass, reach grams. Liters become mL for the % (m/v). ✓

 Practice 6

$0.200 \text{ mol CuSO}_4 = 1 \text{ L of solution}$

the label " 0.200 M CuSO_4 " states this equality · molar mass CuSO_4 159.61 g/mol

A teaching lab fills six volumetric flasks, each to its **250.0 mL** line, with **0.200 M CuSO_4** . What total mass of CuSO_4 , in grams, must be weighed out?

- A. 7.98
- B. 1.20×10^3
- C. 47.9
- D. 0.300
- E. 4.79×10^4

✓ Practice 6: answer C

0.200 mol CuSO₄ = 1 L of solution

given: 6 × 250.0 mL = 1500. mL = 1.500 L of 0.200 M CuSO₄ (159.61 g/mol) · wanted: total g CuSO₄

$$1.500 \text{ L soln} \times \frac{0.200 \text{ mol CuSO}_4}{1 \text{ L soln}} \times \frac{159.61 \text{ g CuSO}_4}{1 \text{ mol CuSO}_4} = 47.9 \text{ g (answer C)}$$

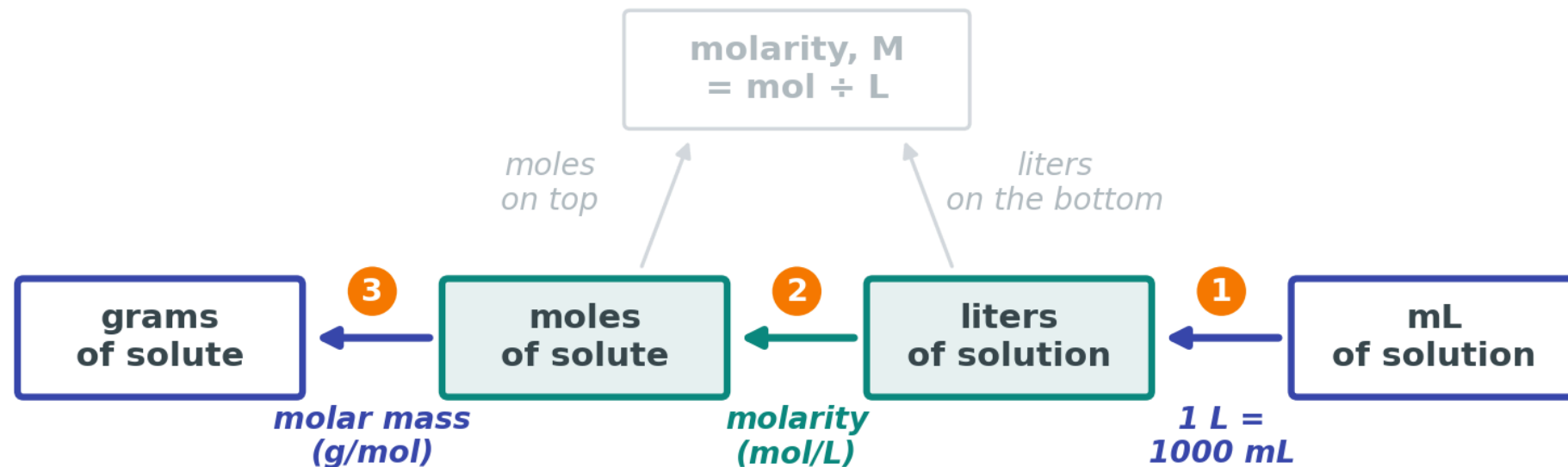
A filled one flask: $0.2500 \times 0.200 \times 159.61 = 7.98$. B flipped the molarity: $1.500 \div 0.200 \times 159.61 = 1.20 \times 10^3$. D stopped at moles: $1.500 \times 0.200 = 0.300 \text{ mol}$. E skipped mL → L: $1500. \times 0.200 \times 159.61 = 4.79 \times 10^4$.

🧠 **Sense check:** Six quarter liters make 1.500 L. That holds 0.300 mol, at about 160 g per mole: about 48 g. 47.9 g ✓

✓ Practice 6: the route on the map

0.200 mol CuSO_4 = 1 L of solution

given: six flasks of 250.0 mL · found: 47.9 g CuSO_4



🧠 **Sense check:** Add the flask volumes first, 1500. mL, then make the three moves of a single flask. ✓

Check yourself

1. A bottle is labeled **3.0 M NaOH**. State the equality the label stores, then write the fraction that converts liters of this solution to moles of NaOH.
2. A flask holds 0.20 mol of solute, with water to the 250.0 mL mark. Which number belongs under the moles in $M = \text{mol} \div L$: 250 or 0.2500?

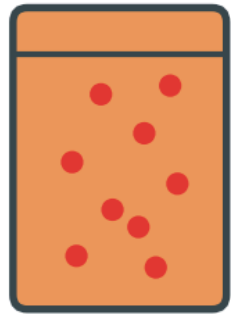
A stock solution is often more concentrated than a procedure needs. Adding water lowers the molarity while the moles of solute stay the same, so $M_1V_1 = M_2V_2$.

6 • Dilution

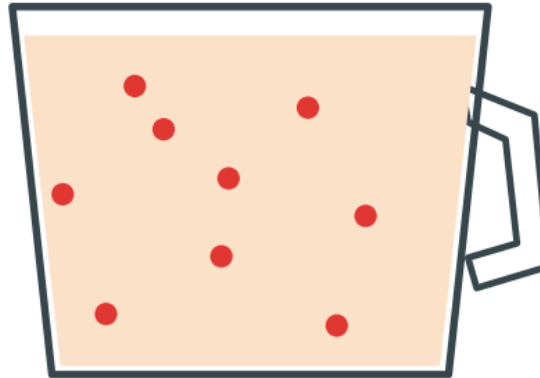
Use $M_1V_1 = M_2V_2$ to find a diluted concentration or the stock volume needed for a target, remembering that adding water conserves the moles of solute and that both volumes must share a unit.

One can makes a whole pitcher

the same flavor bits — spread through much more liquid



+ a pitcher of water



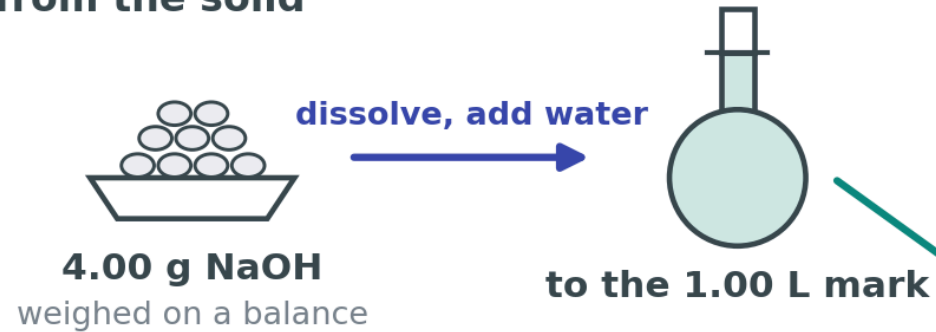
one small can of
thick, strong concentrate

a full pitcher of
pale, mild juice

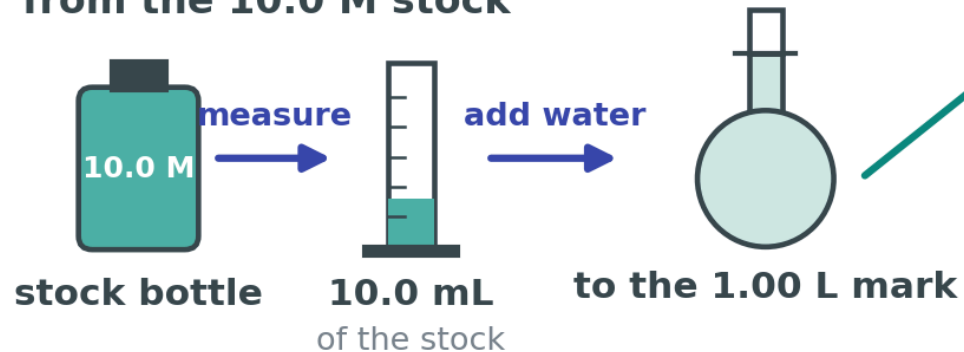
Frozen juice concentrate is thick and strong. Stir one small can into a pitcher of water and it becomes a full, mild drink. Same juice, just more liquid.

A stock bottle skips the weighing

from the solid



from the 10.0 M stock



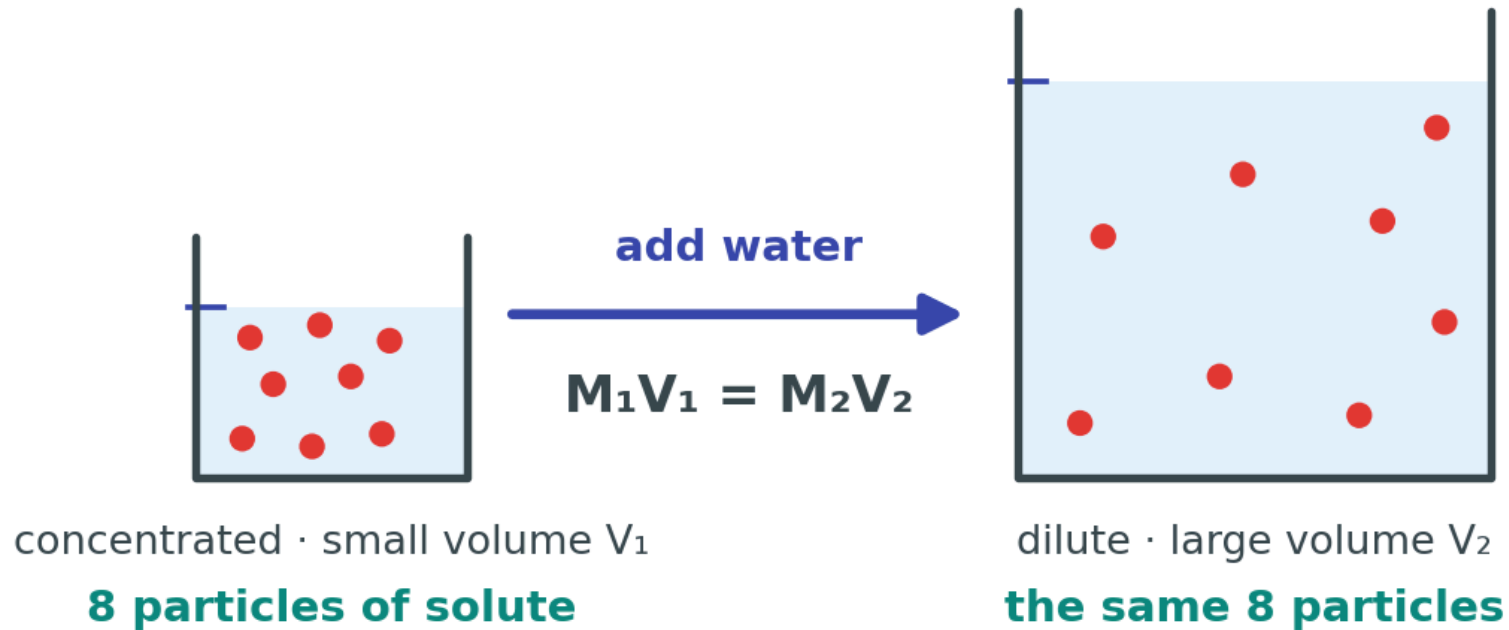
the same flask

1.00 L of 0.100 M NaOH
0.100 mol NaOH either way

NaOH comes as a solid or a 10.0 M stock. Both fill the flask with 0.100 mol. Measuring some stock and adding water is a **dilution**: a weaker solution from a stronger one.

The solute stays; only the water grows

the solute is conserved — only the volume grows



Adding water spreads the solute through more liquid. Not one particle of solute is added or removed, so the moles of solute stay fixed. That conservation is the whole rule: $M_1V_1 = M_2V_2$.

The four quantities in $M_1V_1 = M_2V_2$

both products equal the moles of solute — unchanged by adding water

$$\boxed{M_1} \times \boxed{V_1} = \boxed{M_2} \times \boxed{V_2}$$

concentration before volume before concentration after volume after
mol/L **mL or L** **mol/L** **mL or L**

Before diluting: concentration M_1 and volume V_1 . After: concentration M_2 and volume V_2 . Their products are equal because the moles of solute, concentration times volume, never change.

memory hook: 1 = before, 2 = after

$M \times V$ counts the moles of solute: the same number on both sides

Both volumes in the same unit

$$M_1V_1 = M_2V_2$$

V_1 and V_2 both in mL, or both in L · the volume unit cancels across the equation

The relation balances only when V_1 and V_2 carry the same volume unit. Put both in milliliters or both in liters. The solved volume comes out in whatever unit you used.

$C_1V_1 = C_2V_2$: any concentration unit

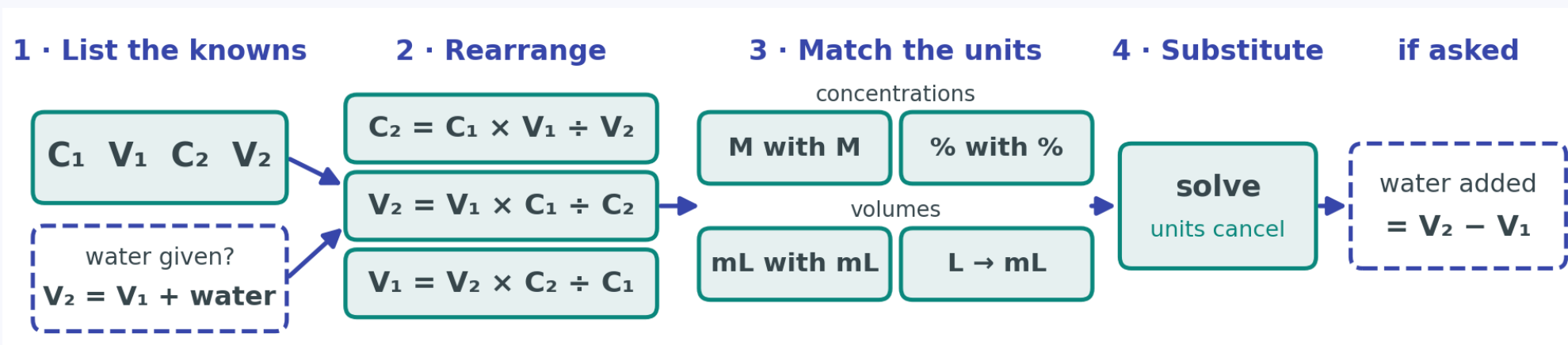
$$C_1V_1 = C_2V_2$$

C in M: $M \times L$ counts moles of solute · C in % (m/v): $\% \times mL \div 100$ counts grams of solute

Adding water changes neither the moles nor the grams of solute. The relation holds for molarity and percent (m/v) alike. C_1 and C_2 share one unit: M with M, % with %.

The method

1. **List the knowns:** three of M_1 , V_1 , M_2 , V_2 ; mark the unknown.
2. **Rearrange for the unknown** in $M_1V_1 = M_2V_2$.
3. **Match the units:** C_1 with C_2 , V_1 with V_2 .
4. **Substitute and solve;** the diluted concentration comes out lower.





Guided example: NaOH from the 10.0 M stock

$$M_1V_1 = M_2V_2$$

given: stock 10.0 M NaOH · target 0.200 M · final volume 250. mL · wanted: mL of stock

A lab needs **250. mL of 0.200 M NaOH**. The shelf holds the **10.0 M NaOH** stock. What volume of the stock, in mL, is measured out?

The stock bottle is "before". The finished flask is "after".

✓ Guided example: solution

$$M_1V_1 = M_2V_2$$

given: stock 10.0 M NaOH · target 0.200 M · final volume 250. mL · wanted: mL of stock

Step 1 · List the knowns

Step 2 · Rearrange

$$M_1V_1 = M_2V_2 \rightarrow V_1 = V_2 \times \frac{M_2}{M_1}$$

Step 3 · Match the units

Step 4 · Substitute and solve

$$V_1 = 250. \text{ mL} \times \frac{0.200 \text{ M}}{10.0 \text{ M}} = 5.00 \text{ mL}$$

 **Sense check:** The stock is 50 times stronger, 10.0 M against 0.200 M, so it fills a fiftieth of the flask: $250. \div 50 = 5.00 \text{ mL}$ ✓

✓ Guided example: the route on the map

$$M_1V_1 = M_2V_2$$

given: 10.0 M stock · 0.200 M target · 250. mL final · found: $V_1 = 5.00$ mL of stock

1 · List the knowns

C_1 V_1 C_2 V_2

water given?
 $V_2 = V_1 + \text{water}$

2 · Rearrange

$$C_2 = C_1 \times V_1 \div V_2$$

$$V_2 = V_1 \times C_1 \div C_2$$

$$V_1 = V_2 \times C_2 \div C_1$$

3 · Match the units

concentrations

M with M

% with %

volumes

mL with mL

L → mL

4 · Substitute

solve
units cancel

if asked

water added
 $= V_2 - V_1$

 **Sense check:** The unknown is the stock volume, so $V_1 = V_2 \times M_2 \div M_1$. Both volumes are in mL, and no water amount was asked. ✓

 Practice 1

$$M_1V_1 = M_2V_2$$

stock 12.0 M HCl · target 0.600 M · batch 300. mL

A titration calls for 0.600 M HCl. How many milliliters of **12.0 M HCl** stock go into a **300. mL** batch?

- A. 6.00×10^3
- B. 180.
- C. 285
- D. 15.0

✓ Practice 1 answer: D

$$M_1V_1 = M_2V_2$$

given: stock 12.0 M · target 0.600 M · 300. mL batch · wanted: mL of stock

$$V_1 = 300. \text{ mL} \times \frac{0.600 \text{ M}}{12.0 \text{ M}} = 15.0 \text{ mL} \cdot \text{answer D}$$

A flipped the ratio: $300. \times (12.0 \div 0.600) = 6.00 \times 10^3 \text{ mL}$, twenty times the batch. B stopped at the solute: $300. \times 0.600 = 180. \text{ mmol}$ of HCl, not a volume. C reported the water: $300. - 15.0 = 285 \text{ mL}$ goes in after the stock.

 **Sense check:** The stock is 20 times stronger, so it fills a twentieth of the batch: $300. \div 20 = 15.0 \text{ mL}$ ✓

✓ Practice 1: the route on the map

$$M_1V_1 = M_2V_2$$

given: stock 12.0 M · target 0.600 M · 300. mL batch · found: $V_1 = 15.0$ mL of stock

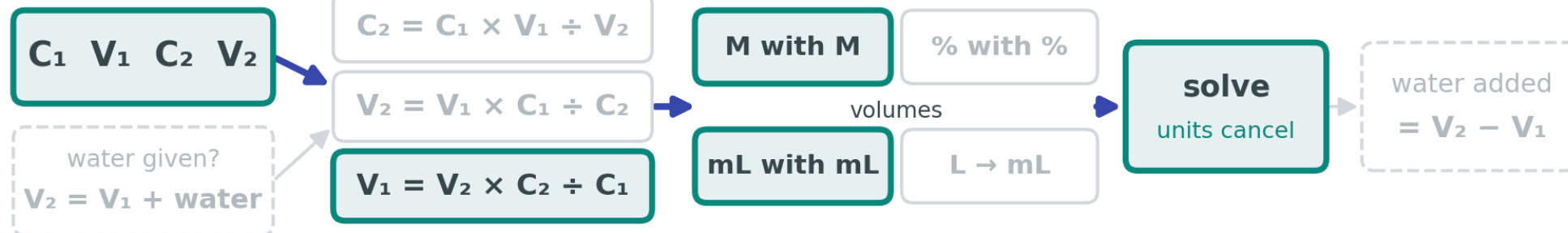
1 · List the knowns

2 · Rearrange

3 · Match the units

4 · Substitute

if asked



Sense check: The unknown is the stock volume, the path of the guided example: $V_1 = V_2 \times M_2 \div M_1$. Both volumes are in mL. ✓



Worked example 1: concentration after dilution

$$M_1V_1 = M_2V_2$$

given: 6.00 M · 25.0 mL stock · water to 150. mL final · wanted: M_2

A stockroom takes **25.0 mL of 6.00 M HCl** and adds water to a final volume of **150. mL**. Find the concentration of the diluted solution.

List the knowns, then rearrange $M_1V_1 = M_2V_2$ for M_2 .

✓ Worked example 1: solution

$$M_1V_1 = M_2V_2$$

given: 6.00 M · 25.0 mL stock · water to 150. mL final · wanted: M_2

Step 1 · List the knowns

Step 2 · Rearrange

$$M_1V_1 = M_2V_2 \rightarrow M_2 = M_1 \times \frac{V_1}{V_2}$$

Step 3 · Match the units

Step 4 · Substitute and solve

$$M_2 = 6.00 \text{ M} \times \frac{25.0 \text{ mL}}{150. \text{ mL}} = 1.00 \text{ M}$$

 **Sense check:** The volume grew six-fold, from 25.0 to 150. mL, so the concentration falls six-fold: $6.00 \div 6 = 1.00 \text{ M}$.
Diluting lowers the concentration ✓

✓ Worked example 1: the route on the map

$$M_1V_1 = M_2V_2$$

given: 6.00 M · 25.0 mL stock · 150. mL final · found: $M_2 = 1.00$ M

1 · List the knowns

C_1 V_1 C_2 V_2

water given?
 $V_2 = V_1 + \text{water}$

2 · Rearrange

$$C_2 = C_1 \times V_1 \div V_2$$

$$V_2 = V_1 \times C_1 \div C_2$$

$$V_1 = V_2 \times C_2 \div C_1$$

3 · Match the units

concentrations

M with M

% with %

volumes

mL with mL

L → mL

4 · Substitute

solve
units cancel

if asked

water added
 $= V_2 - V_1$

 **Sense check:** The unknown is the new concentration, so $M_2 = M_1 \times V_1 \div V_2$. Both volumes are in mL, so their ratio is a pure number. ✓



Worked example 2: final volume after dilution

$$M_1V_1 = M_2V_2$$

given: 6.00 M · 30.0 mL stock · dilute to 0.500 M · wanted: V_2

How much dilute solution results when **30.0 mL of 6.00 M NaOH** is diluted with water to **0.500 M**? Find the total volume.

A tempting setup puts the lower concentration on top. Weigh it against the sense check.

✓ Worked example 2: solution

$$M_1V_1 = M_2V_2$$

given: 6.00 M · 30.0 mL stock · dilute to 0.500 M · wanted: V_2

A common first attempt

$$V_2 = 30.0 \text{ mL} \times \frac{0.500 \text{ M}}{6.00 \text{ M}} = 2.50 \text{ mL} \times$$

Step 1 · List the knowns

Step 2 · Rearrange

Step 3 · Match the units

Step 4 · Substitute and solve

$$V_2 = 30.0 \text{ mL} \times \frac{6.00 \text{ M}}{0.500 \text{ M}} = \mathbf{360 \text{ mL}}$$

🧠 **Sense check:** 6.00 M down to 0.500 M is a twelve-fold drop, so the volume grows: $30.0 \text{ mL} \times 12 = 360 \text{ mL}$ ✓

✓ Worked example 2: the route on the map

$$M_1V_1 = M_2V_2$$

given: 6.00 M · 30.0 mL stock · dilute to 0.500 M · found: $V_2 = 360$ mL

1 · List the knowns

C_1 V_1 C_2 V_2

water given?
 $V_2 = V_1 + \text{water}$

2 · Rearrange

$$C_2 = C_1 \times V_1 \div V_2$$

$$V_2 = V_1 \times C_1 \div C_2$$

$$V_1 = V_2 \times C_2 \div C_1$$

3 · Match the units

concentrations

M with M

% with %

volumes

mL with mL

L → mL

4 · Substitute

solve
units cancel

if asked

water added
 $= V_2 - V_1$

 **Sense check:** The unknown is the final volume, so $V_2 = V_1 \times M_1 \div M_2$. The question asked for the total volume, so the water step stays unlit. ✓

Take-home: diluting grows the volume

$$V_2 = 30.0 \text{ mL} \times \frac{0.500 \text{ M}}{6.00 \text{ M}} = 2.50 \text{ mL} \times \cdot \text{smaller than the start}$$

$$V_2 = 30.0 \text{ mL} \times \frac{6.00 \text{ M}}{0.500 \text{ M}} = \mathbf{360 \text{ mL} \checkmark}$$

Diluting spreads the solute through more liquid, so the final volume exceeds the start. Keep the higher starting concentration on top, and the volume grows. A shrinking result means the ratio was flipped.

Your turn: glucose from a stock

$$M_1V_1 = M_2V_2$$

given: stock 1.50 M glucose · dilute to 0.300 M · final volume 250. mL · wanted: V_1

$$V_1 = V_2 \times \frac{M_2}{M_1} = \text{_____ mL} \times \frac{\text{_____ M}}{\text{_____ M}} = \text{_____ mL}$$

$$V_1 = V_2 \times \frac{M_2}{M_1} = 250. \text{ mL} \times \frac{0.300 \text{ M}}{1.50 \text{ M}} = \mathbf{50.0 \text{ mL}}$$

 **Sense check:** The stock is five times stronger, 1.50 M against 0.300 M, so it takes a fifth of the final volume: $250. \div 5 = 50.0 \text{ mL} \checkmark$

⚠ Where this goes wrong

6.00 M · 30.0 mL stock · dilute to 0.500 M

correct: $V_2 = 30.0 \text{ mL} \times (6.00 \text{ M} \div 0.500 \text{ M}) = 360 \text{ mL}$

Flipping the concentration ratio. $30.0 \text{ mL} \times (0.500 \text{ M} \div 6.00 \text{ M}) = 2.50 \text{ mL}$ claims diluting shrank the liquid. The higher starting concentration goes on top: $30.0 \times (6.00 \div 0.500) = 360 \text{ mL}$.

Stopping at the solute amount. $30.0 \text{ mL} \times 6.00 \text{ M} = 180$ gives the millimoles of solute, not a volume. Divide that by the new concentration: $180 \div 0.500 = 360 \text{ mL}$.

Reporting only the water added. $360 - 30.0 = 330 \text{ mL}$ is the water poured in. The total volume still holds the original 30.0 mL of stock: 360 mL.



Practice 2

$$M_1V_1 = M_2V_2$$

start 5.00 M HNO_3 · dilute to 0.400 M

You have **40.0 mL of 5.00 M nitric acid (HNO_3)** and dilute it with water until the concentration is **0.400 M**. What volume of water, in mL, must be added?

- A. 500.
- B. 460.
- C. 3.20
- D. 200.

✓ Practice 2 answer: B

$$M_1V_1 = M_2V_2$$

given: 5.00 M · 40.0 mL HNO₃ · dilute to 0.400 M · wanted: mL of water added

$$V_2 = 40.0 \text{ mL} \times \frac{5.00 \text{ M}}{0.400 \text{ M}} = 500. \text{ mL total} \rightarrow \text{water} = 500. - 40.0 = \mathbf{460. \text{ mL} \cdot \text{answer B}}$$

A stopped at the total volume: 500. mL still counts the 40.0 mL of acid already in the flask. C flipped the ratio: $40.0 \times (0.400 \div 5.00) = 3.20 \text{ mL}$, less than the start. D stopped at the solute: $40.0 \times 5.00 = 200. \text{ mmol}$ of HNO₃, not a volume.

 **Sense check:** 5.00 M down to 0.400 M is a 12.5-fold drop, so the solution grows to 500. mL; the acid supplied 40.0 mL of it and water the other 460. mL ✓

✓ Practice 2: the route on the map

$$M_1V_1 = M_2V_2$$

given: 5.00 M · 40.0 mL HNO₃ · dilute to 0.400 M · found: 460. mL of water added

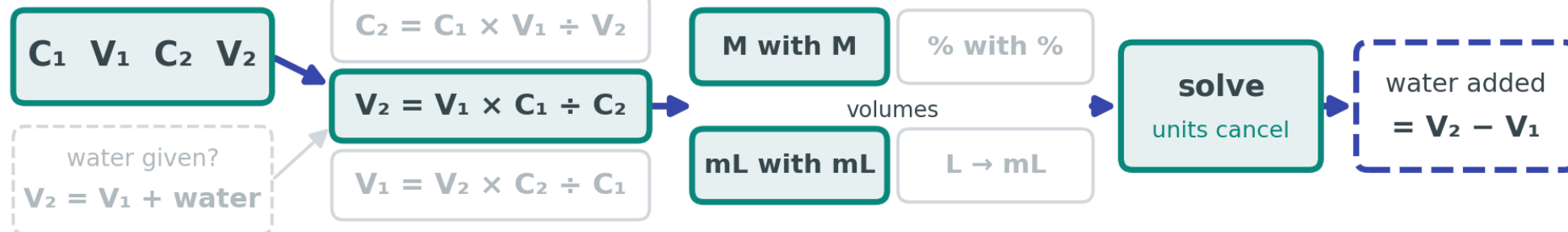
1 · List the knowns

2 · Rearrange

3 · Match the units

4 · Substitute

if asked



🧠 **Sense check:** $V_2 = V_1 \times M_1 \div M_2$ gives the total, 500. mL. The question asked for the water, so the last step subtracts the 40.0 mL of acid. ✓



Worked example 3: a percent dilution

$$C_1V_1 = C_2V_2$$

given: 9.00% (m/v) NaOH · 10.0 mL stock · water to 60.0 mL final · wanted: % (m/v) after

A technician dilutes **10.0 mL of 9.00% (m/v) NaOH** with water to a final volume of **60.0 mL**. Find the new percent (m/v).

Percent is the concentration unit here. The steps do not change.

✓ Worked example 3: solution

$$C_1V_1 = C_2V_2$$

given: 9.00% (m/v) · 10.0 mL stock · water to 60.0 mL final · wanted: C_2 in % (m/v)

A common first attempt

$$C_2 = 9.00\% \times \frac{10.0 \text{ mL}}{50.0 \text{ mL water}} = 1.80\% \times$$

Step 1 · List the knowns

Step 2 · Rearrange

Step 3 · Match the units

Step 4 · Substitute and solve

$$C_2 = 9.00\% \times \frac{10.0 \text{ mL}}{60.0 \text{ mL}} = 1.50\% \text{ (m/v)}$$

 **Sense check:** Grams check: $10.0 \text{ mL} \times 9.00 \text{ g/100 mL} = 0.900 \text{ g NaOH}$, and 0.900 g in 60.0 mL is $1.50 \text{ g per 100 mL}$ ✓

✓ Worked example 3: the route on the map

$$C_1V_1 = C_2V_2$$

given: 9.00% (m/v) · 10.0 mL stock · 60.0 mL final · found: $C_2 = 1.50\%$ (m/v)

1 · List the knowns

C_1 V_1 C_2 V_2

water given?
 $V_2 = V_1 + \text{water}$

2 · Rearrange

$$C_2 = C_1 \times V_1 \div V_2$$

$$V_2 = V_1 \times C_1 \div C_2$$

$$V_1 = V_2 \times C_2 \div C_1$$

3 · Match the units

concentrations

M with M

% with %

volumes

mL with mL

L → mL

4 · Substitute

solve
units cancel

if asked

water added
 $= V_2 - V_1$

🧠 **Sense check:** The same path as worked example 1. Only the concentration unit changed: % with %. ✓



Practice 3

$$C_1V_1 = C_2V_2$$

start 8.00% (m/v) KCl · 15.0 mL · plus 25.0 mL of water

A pharmacist mixes **15.0 mL of 8.00% (m/v) KCl** with **25.0 mL of water**. What is the new concentration, in % (m/v)?

- A. 3.00
- B. 4.80
- C. 21.3
- D. 1.20

✓ Practice 3 answer: A

$$C_1V_1 = C_2V_2$$

given: 8.00% (m/v) · 15.0 mL stock · 25.0 mL water added · wanted: C_2 in % (m/v)

$$V_2 = 15.0 \text{ mL} + 25.0 \text{ mL} = 40.0 \text{ mL} \rightarrow C_2 = 8.00\% \times \frac{15.0 \text{ mL}}{40.0 \text{ mL}} = 3.00\% \cdot \text{answer A}$$

B divided by the water alone, per solvent: $8.00 \times (15.0 \div 25.0) = 4.80$. C flipped the volume ratio: $8.00 \times (40.0 \div 15.0) = 21.3$, stronger than the start. D stopped at the solute: $15.0 \text{ mL} \times 8.00 \text{ g/100 mL} = 1.20 \text{ g}$ of KCl, not a percent.

🧠 **Sense check:** The 1.20 g of KCl now sits in 40.0 mL of solution: $1.20 \text{ g} \div 40.0 \text{ mL} \times 100 = 3.00\% \checkmark$

✓ Practice 3: the route on the map

$$C_1V_1 = C_2V_2$$

given: 8.00% (m/v) · 15.0 mL stock · 25.0 mL water added · found: $C_2 = 3.00\%$ (m/v)

1 · List the knowns

C_1 V_1 C_2 V_2

water given?
 $V_2 = V_1 + \text{water}$

2 · Rearrange

$$C_2 = C_1 \times V_1 \div V_2$$

$$V_2 = V_1 \times C_1 \div C_2$$

$$V_1 = V_2 \times C_2 \div C_1$$

3 · Match the units

concentrations

M with M

% with %

volumes

mL with mL

L → mL

4 · Substitute

solve
units cancel

if asked

water added
 $= V_2 - V_1$

🧠 **Sense check:** The water was given, so $V_2 = 15.0 + 25.0 = 40.0$ mL first. Then $C_2 = C_1 \times V_1 \div V_2$, the path of worked example 3. ✓



Worked example 4: stock volume for a target

$$M_1V_1 = M_2V_2$$

given: dilute to 0.150 M · final volume 2.00 L · stock 6.00 M · wanted: mL of stock

A titration needs **2.00 L of 0.150 M HCl**. The stockroom stocks **6.00 M HCl**. What volume of the stock, in milliliters, do you measure out?

Track the unit on every volume as you go.

✓ Worked example 4: solution

$$M_1V_1 = M_2V_2$$

given: dilute to 0.150 M · final volume 2.00 L · stock 6.00 M · wanted: mL of stock

A common first attempt

$$V_1 = 2.00 \text{ L} \times \frac{0.150 \text{ M}}{6.00 \text{ M}} = 0.0500 \text{ mL} \times$$

Step 1 · List the knowns

Step 2 · Rearrange

Step 3 · Match the units

Step 4 · Substitute and solve

$$V_1 = 2.00 \text{ L} \times \frac{0.150 \text{ M}}{6.00 \text{ M}} = 0.0500 \text{ L} = \mathbf{50.0 \text{ mL}}$$

🧠 **Sense check:** The stock is forty times stronger than 0.150 M, so V_1 is a fortieth of 2.00 L: 0.0500 L, 50.0 mL ✓

✓ Worked example 4: the route on the map

$$M_1V_1 = M_2V_2$$

given: 0.150 M · 2.00 L final · 6.00 M stock · found: $V_1 = 50.0$ mL

1 · List the knowns

C_1 V_1 C_2 V_2

water given?
 $V_2 = V_1 + \text{water}$

2 · Rearrange

$$C_2 = C_1 \times V_1 \div V_2$$

$$V_2 = V_1 \times C_1 \div C_2$$

$$V_1 = V_2 \times C_2 \div C_1$$

3 · Match the units

concentrations

M with M

% with %

volumes

mL with mL

L → mL

4 · Substitute

solve
units cancel

if asked

water added
 $= V_2 - V_1$

 **Sense check:** The unknown is the stock volume, as in the guided example. The new move is the unit: V_2 in liters, the answer wanted in mL. ✓



Practice 4

$$M_1V_1 = M_2V_2$$

stock 4.50 M KOH · target 0.400 M · final volume 0.900 L

A lab needs **0.900 L of 0.400 M KOH**, made from a **4.50 M KOH** stock. How many milliliters of water are added to the measured stock?

- A. 80.0
- B. 0.360
- C. 820.
- D. 0.820

✓ Practice 4 answer: C

$$M_1V_1 = M_2V_2$$

given: 0.400 M · 0.900 L final · stock 4.50 M · wanted: mL of water added

$$V_1 = 0.900 \text{ L} \times \frac{0.400 \text{ M}}{4.50 \text{ M}} = 0.0800 \text{ L} = 80.0 \text{ mL of stock}$$

$$\text{water} = 900. \text{ mL} - 80.0 \text{ mL} = \mathbf{820. \text{ mL} \cdot \text{answer C}}$$

A stopped at the stock: 80.0 mL is what gets measured, and water fills the rest. B stopped at the solute: $0.900 \times 0.400 = 0.360$ mol of KOH, not a volume. D subtracted in liters: $0.900 - 0.0800 = 0.820$ L, not mL; convert both volumes first.

 **Sense check:** The stock is 11.25 times stronger than the target, so it fills only a small share of the 900. mL; water fills the other 820. mL ✓

✓ Practice 4: the route on the map

$$M_1V_1 = M_2V_2$$

given: 0.400 M · 0.900 L final · stock 4.50 M · found: 820. mL of water added

1 · List the knowns

2 · Rearrange

3 · Match the units

4 · Substitute

if asked



Sense check: V_2 entered in liters, so 0.900 L becomes 900. mL before the subtraction. Stock first, 80.0 mL, then water = 900. – 80.0. ✓



Practice 5

$$C_1V_1 = C_2V_2$$

stock 18.0% (m/v) NaCl · target 0.900% (m/v) · final volume 3.00 L

A pharmacy prepares **3.00 L of 0.900% (m/v) saline** from an **18.0% (m/v) NaCl** stock. What volume of water, in mL, goes in with the measured stock?

- A. 150.
- B. 3.00×10^3
- C. 2.85
- D. 27.0
- E. 2.85×10^3

✓ Practice 5 answer: E

$$C_1V_1 = C_2V_2$$

given: $0.900\% \text{ (m/v)} \cdot 3.00 \text{ L} = 3.00 \times 10^3 \text{ mL final} \cdot 18.0\% \text{ (m/v) stock} \cdot \text{wanted: mL of water}$

$$V_1 = 3.00 \times 10^3 \text{ mL} \times \frac{0.900\%}{18.0\%} = 150. \text{ mL of stock}$$

$$\text{water} = 3.00 \times 10^3 \text{ mL} - 150. \text{ mL} = \mathbf{2.85 \times 10^3 \text{ mL} \cdot \text{answer E}}$$

A stopped at the stock: 150. mL is what gets measured. B read 0.900% per solvent, 27.0 g NaCl per $3.00 \times 10^3 \text{ mL}$ of water; the stock fills part of the 3.00 L. C subtracted in liters: $3.00 - 0.150 = 2.85 \text{ L}$, not mL. D stopped at the solute: $3.00 \times 10^3 \text{ mL} \times 0.900 \text{ g/100 mL} = 27.0 \text{ g}$ of NaCl.

 **Sense check:** The stock is 20 times stronger, so it fills a twentieth of the 3.00 L; water fills the rest ✓

✓ Practice 5: the route on the map

$$C_1V_1 = C_2V_2$$

given: 0.900% (m/v) · 3.00 L final · 18.0% (m/v) stock · found: 2.85×10^3 mL of water

1 · List the knowns

2 · Rearrange

3 · Match the units

4 · Substitute

if asked



🧠 **Sense check:** Concentrations pair % with %, volumes need L → mL: 3.00 L becomes 3.00×10^3 mL, the stock is 150. mL, and water fills the rest. ✓



Extra practice 1

$$M_1V_1 = M_2V_2$$

start 2.40 M NaCl · 35.0 mL · plus 85.0 mL of water

A lab aide mixes **35.0 mL of 2.40 M NaCl** with **85.0 mL of water**. What is the molarity of the new solution?

- A. 0.700
- B. 0.988
- C. 84.0
- D. 8.23

✓ Extra practice 1 answer: A

$$M_1V_1 = M_2V_2$$

given: 2.40 M NaCl · 35.0 mL stock · 85.0 mL water added · wanted: M_2

$$V_2 = 35.0 \text{ mL} + 85.0 \text{ mL} = 120.0 \text{ mL} \rightarrow M_2 = 2.40 \text{ M} \times \frac{35.0 \text{ mL}}{120.0 \text{ mL}} = \mathbf{0.700 \text{ M} \cdot \text{answer A}}$$

B divided by the water alone: $2.40 \times (35.0 \div 85.0) = 0.988$; the 35.0 mL of stock is part of the solution too. C stopped at the solute: $35.0 \times 2.40 = 84.0$ mmol of NaCl, not a molarity. D flipped the volume ratio: $2.40 \times (120.0 \div 35.0) = 8.23$, stronger than the stock.

 **Sense check:** The volume grew 3.43-fold, 35.0 mL to 120.0 mL, so the molarity falls 3.43-fold: $2.40 \div 3.43 = 0.700$ M ✓



Extra practice 2

$$M_1V_1 = M_2V_2$$

start 4.00 M NaOH · 28.0 mL · final volume 0.350 L

Water is added to **28.0 mL of 4.00 M NaOH** until the volume reaches **0.350 L**. What is the molarity of the diluted solution?

- A. 320.
- B. 50.0
- C. 112
- D. 0.320
- E. 0.296

✓ Extra practice 2 answer: D

$$M_1V_1 = M_2V_2$$

given: 4.00 M NaOH · 28.0 mL stock · 0.350 L = 350. mL final · wanted: M_2

$$M_2 = 4.00 \text{ M} \times \frac{28.0 \text{ mL}}{350. \text{ mL}} = \mathbf{0.320 \text{ M} \cdot \text{answer D}}$$

A mixed the units: $4.00 \times (28.0 \div 0.350) = 320.$, eighty times the stock; mL over L does not cancel. B flipped the volume ratio: $4.00 \times (350. \div 28.0) = 50.0$, stronger than the stock. C stopped at the solute: $28.0 \times 4.00 = 112$ mmol of NaOH, not a molarity. E read 0.350 L as the water added: $4.00 \times (28.0 \div 378) = 0.296$; 0.350 L is the final volume, stock included.

 **Sense check:** In mL, the volume grew 12.5-fold, 28.0 to 350., so the molarity falls 12.5-fold: $4.00 \div 12.5 = 0.320 \text{ M}$





Extra practice 3

$$M = \text{mol} \div L, \text{ then } M_1V_1 = M_2V_2$$

stock 8.33 g CaCl_2 in 250.0 mL · a 15.0 mL portion diluted to 100.0 mL

A technician dissolves **8.33 g of CaCl_2** (110.98 g/mol) in water to make **250.0 mL** of stock. A **15.0 mL** portion of the stock is diluted with water to **100.0 mL**. What is the molarity of the final solution?

- A. 0.300
- B. 5.00
- C. 0.0751
- D. 2.00
- E. 0.0450

✓ Extra practice 3 answer: E

$$M = \text{mol} \div L, \text{ then } M_1V_1 = M_2V_2$$

given: 8.33 g CaCl_2 (110.98 g/mol) in 250.0 mL · 15.0 mL portion diluted to 100.0 mL · wanted: M_2

$$8.33 \text{ g} \div 110.98 \text{ g/mol} = 0.0751 \text{ mol} \rightarrow M_1 = 0.0751 \text{ mol} \div 0.2500 \text{ L} = 0.300 \text{ M}$$

$$M_2 = 0.300 \text{ M} \times \frac{15.0 \text{ mL}}{100.0 \text{ mL}} = \mathbf{0.0450 \text{ M} \cdot \text{answer E}}$$

A stopped at the stock: 0.300 M, before the dilution. B divided grams by liters: $8.33 \div 0.2500 = 33.3 \text{ g/L}$, diluted to 5.00, not a molarity. C stopped at the solute: 0.0751 mol, an amount. D flipped the volume ratio: $0.300 \times (100.0 \div 15.0) = 2.00$, stronger than the stock.

 **Sense check:** 15.0 mL spread to 100.0 mL, 6.67 times the volume: $0.300 \div 6.67 = 0.0450 \text{ M}$ ✓



Check yourself

1. A bottle reads **6.00 M NaOH**. Write **$M_1V_1 = M_2V_2$** for making **250. mL of 0.300 M NaOH**, with the three known values filled in. Which volume is larger: the stock you measure, or 250. mL?
2. You dilute a stock and your result for the final volume comes out smaller than the volume you started with. Name the error, and state which concentration belongs on top of the ratio.

Molarity reports solute per liter of solution. The same amount can be reported per kilogram of solvent (molality), or as a percent by mass or by volume. Diluting still conserves the solute; only the per-amount unit changes.

Can you...?

- ☐ calculate the molar mass of a compound from its formula?
- ☐ convert among grams, moles, and particles with molar mass and Avogadro's number, including atoms or ions of one element inside a compound?
- ☐ calculate the percent composition of a compound from its formula or from measured masses, and use it to find the mass of an element in a sample?
- ☐ reduce a molecular formula to its empirical formula, and find an empirical formula from percent or mass data?
- ☐ scale an empirical formula to the molecular formula with the molar mass?

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