

Moles & Composition

CHEM 50 · Dr. Karmach

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

- Calculate the molar mass of any compound from its formula
- Convert among grams, moles, and particles using molar mass and Avogadro's number
- Calculate the percent composition of a compound from its formula
- Calculate molarity and use it as a conversion factor between solution volume and moles of solute
- Determine empirical and molecular formulas from experimental data
- Express solution concentration as mass, volume, or mass-volume percent and use it as a conversion factor

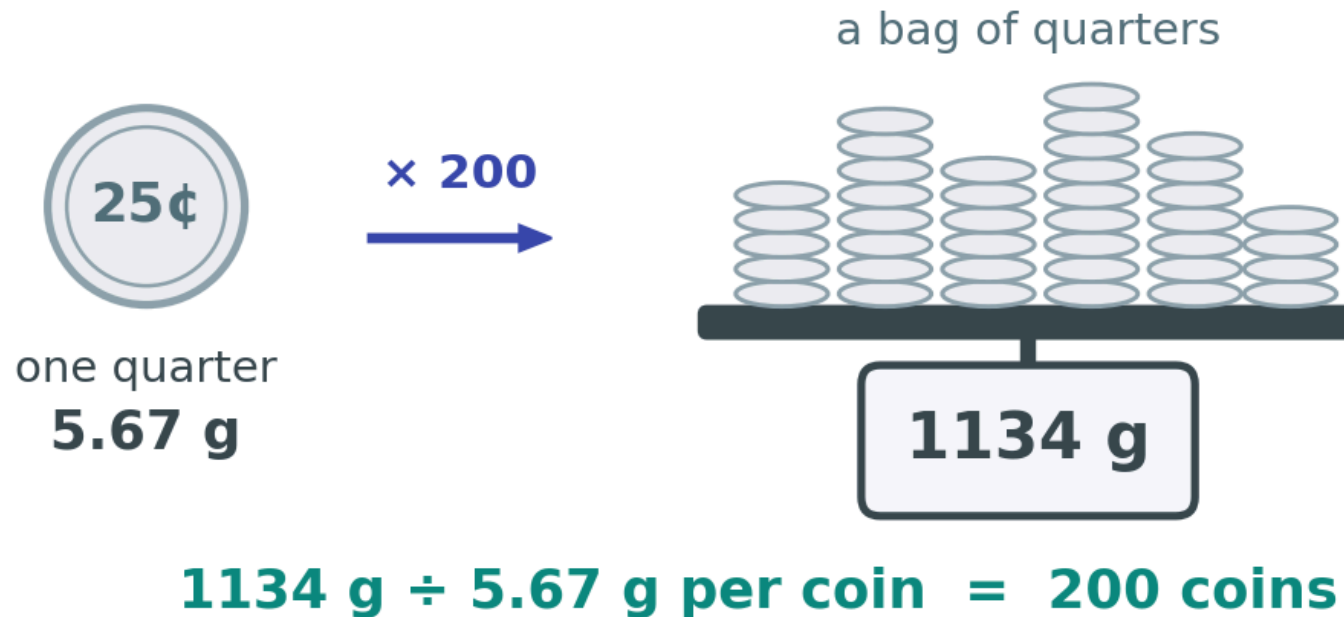
Today's route

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

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.

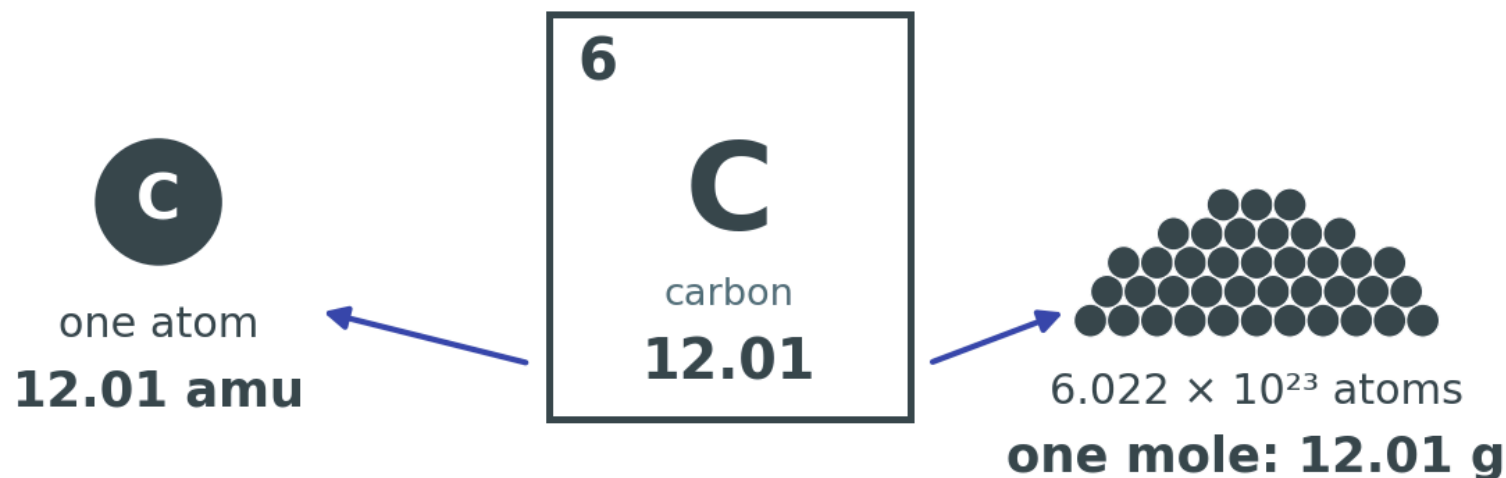
The mole

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

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.

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.

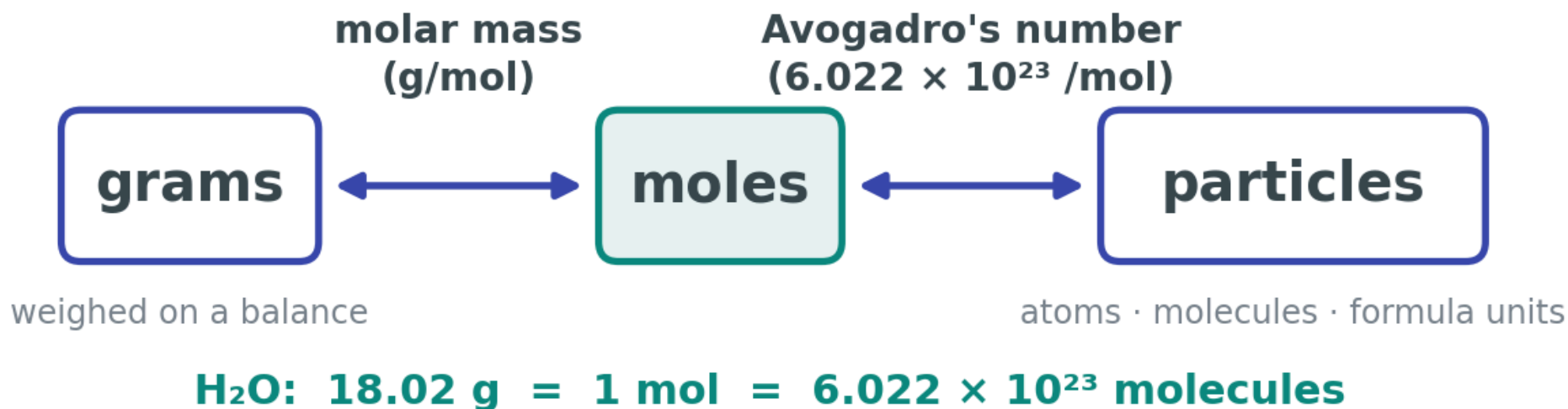
$$\text{H}_2\text{O}: 2(1.008) + 16.00 = 18.02 \text{ g/mol}$$

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.**



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

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

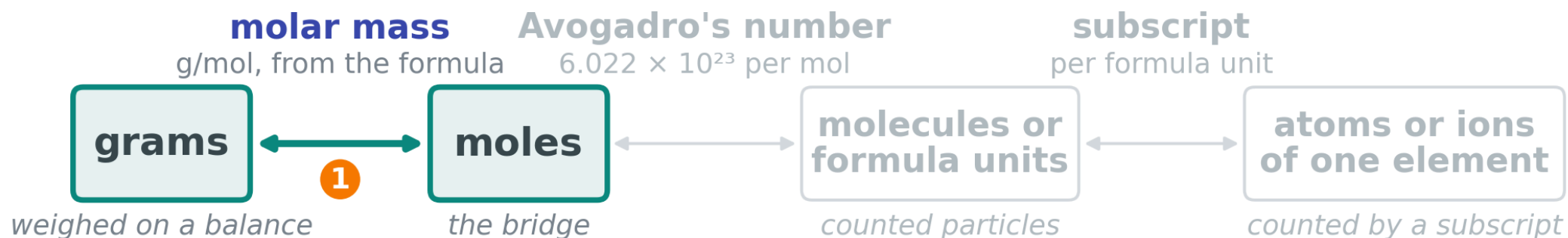
$40.08 + 28.02 + 96.00 \cdot 1 \text{ Ca} \cdot 2 \text{ N} \cdot 6 \text{ O} \cdot \text{every atom in the sum}$

 **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

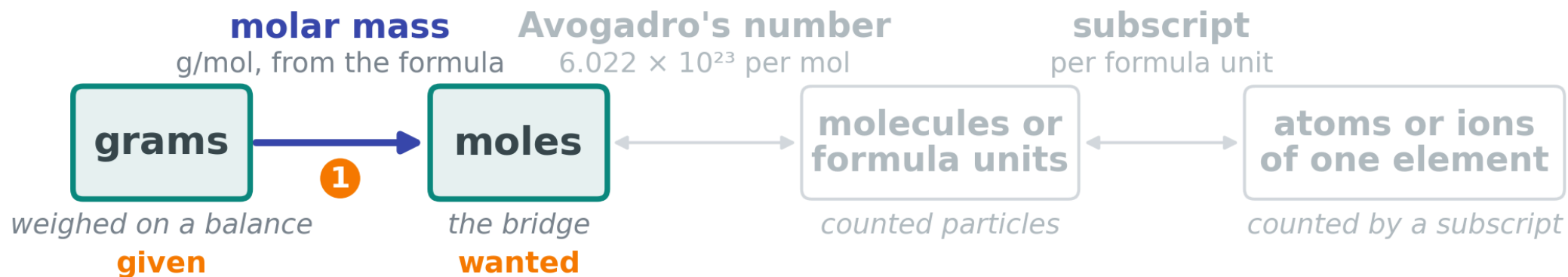
$$25.0 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 0.428 \text{ mol NaCl}$$

 **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. ✓

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.

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

⚠ 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 1

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

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 1: 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

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

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. ✓



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

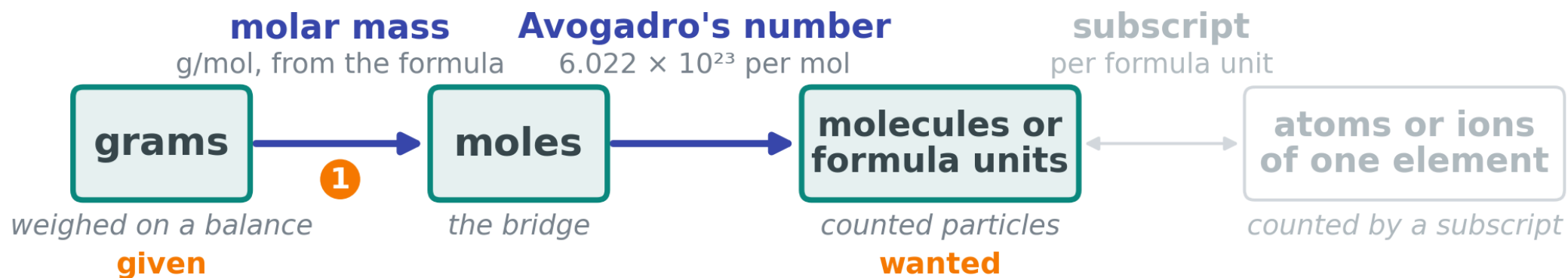
$$0.0500 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol H}_2\text{O}} = 1.67 \times 10^{21} \text{ molecules}$$

 **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 2

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

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 2: 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} . ✓



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

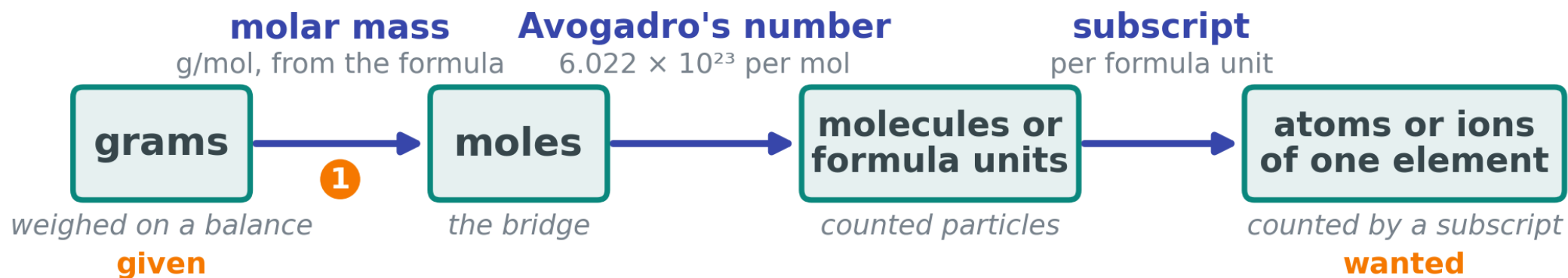
Step 4 · Compute and sense-check

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

 **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 3

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 3: answer D



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

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

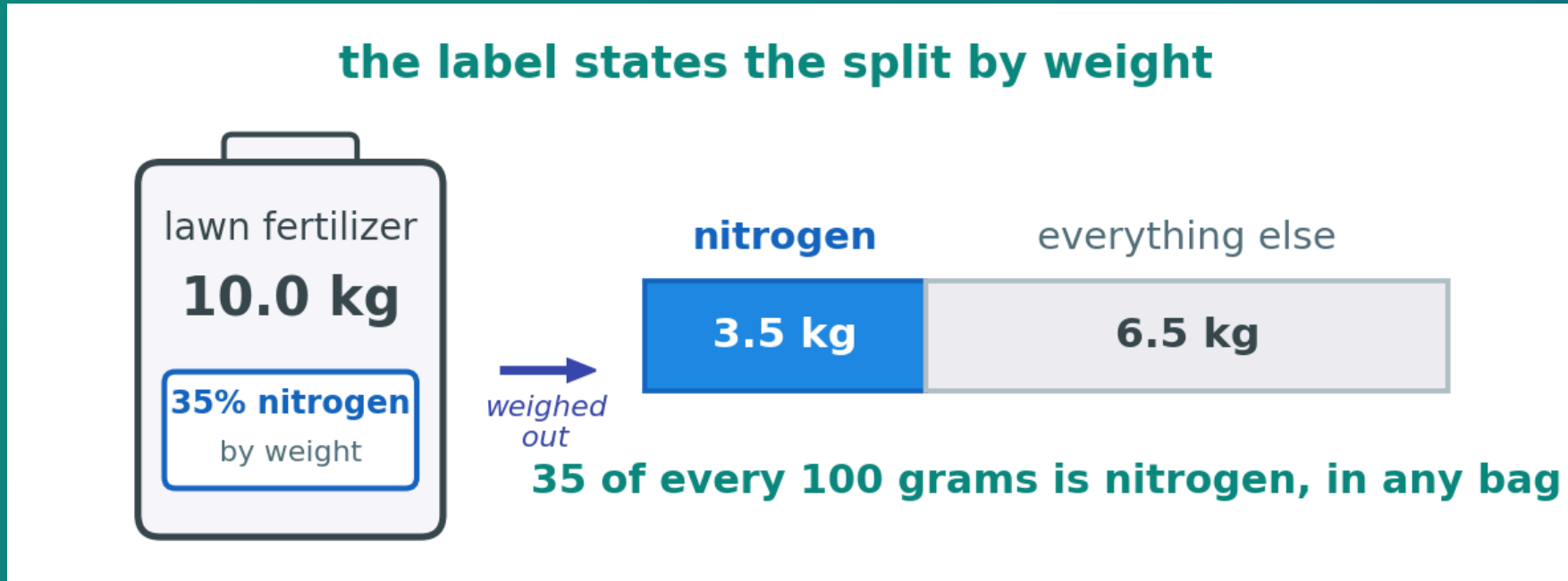
A counted every C atom as a molecule: $1.20 \times 10^{24} \div 6.022 \times 10^{23} = 1.99 \text{ mol}$, $\times 180.16 = 359$. B stopped at moles: 0.332 mol of glucose, never converted to grams. C flipped the subscript factor: $\times 6$ instead of $\div 6$ gives 11.96 mol and $2.15 \times 10^3 \text{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

 **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. ✓

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.

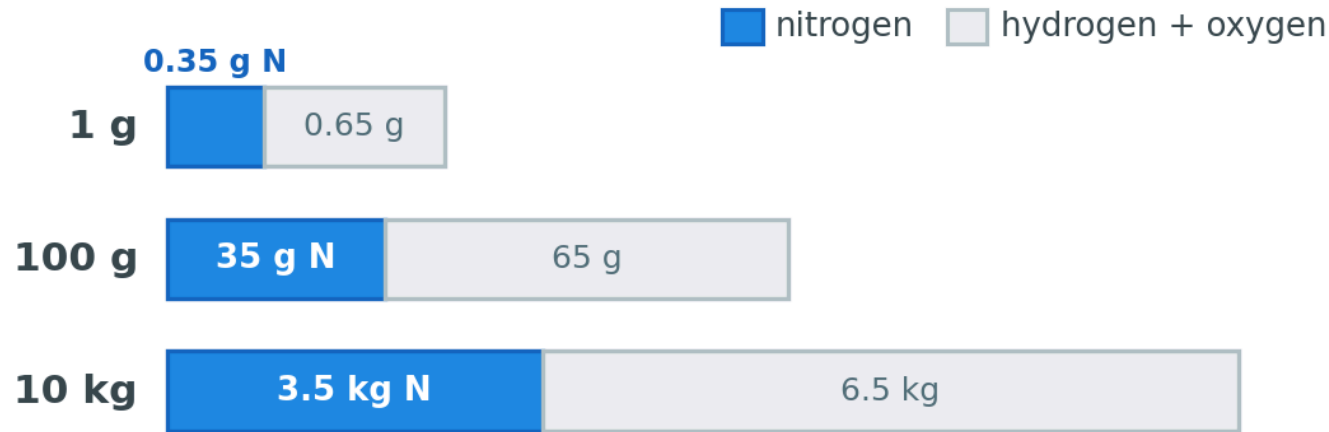
The number on the bag



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.

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

mass % of an element = element mass in one mole ÷ molar mass × 100

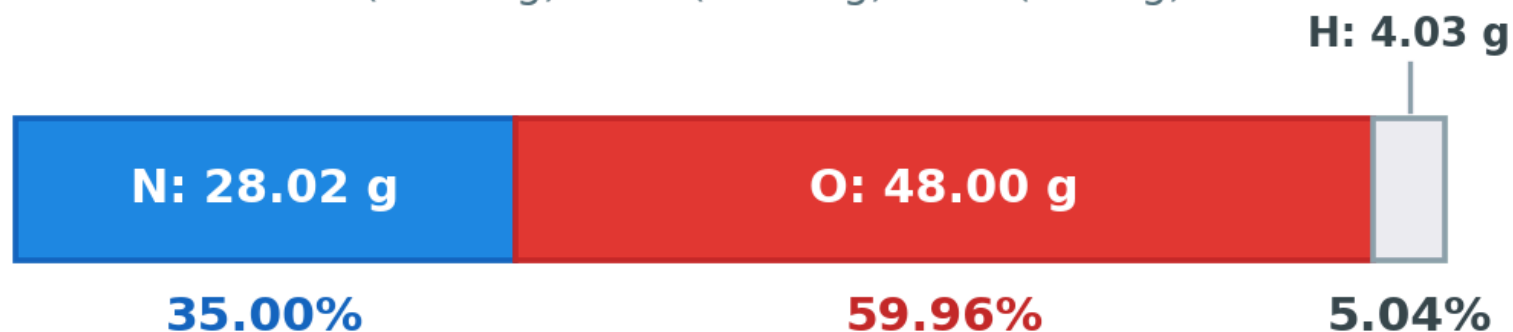
element mass in one mole = atomic mass × subscript

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, × 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, $\times 100$.



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

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

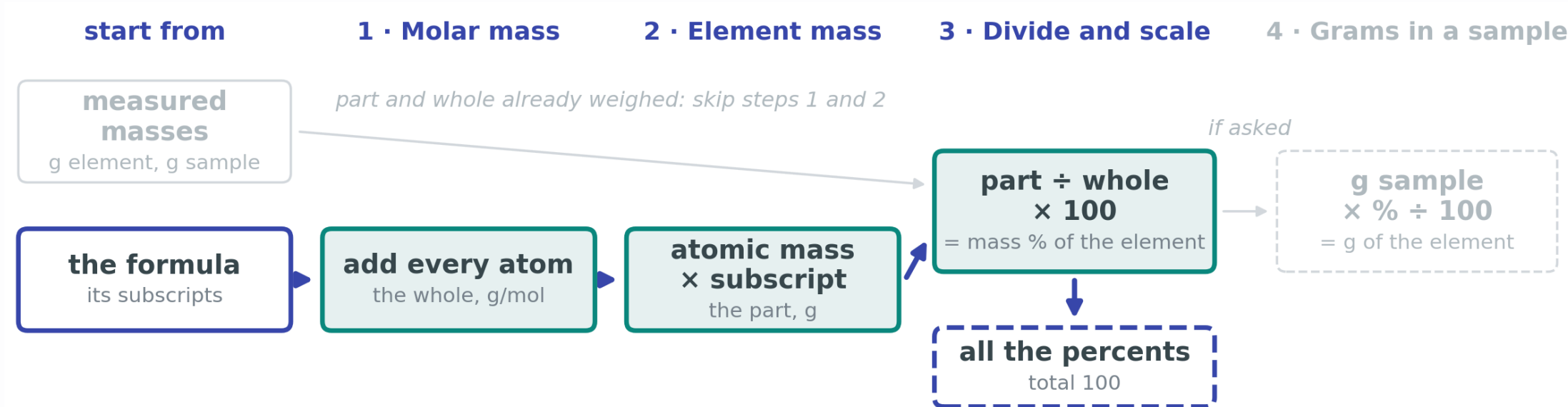
$$\%Cl = \frac{35.45 \text{ g Cl}}{74.55 \text{ g KCl}} \times 100 = \mathbf{47.55\% Cl}$$

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

✓ Worked example 1: the route on the map

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

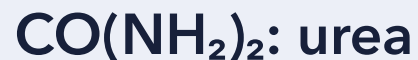
$$\text{found: } 52.45\% \text{ K} \cdot 47.55\% \text{ Cl}$$



 **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

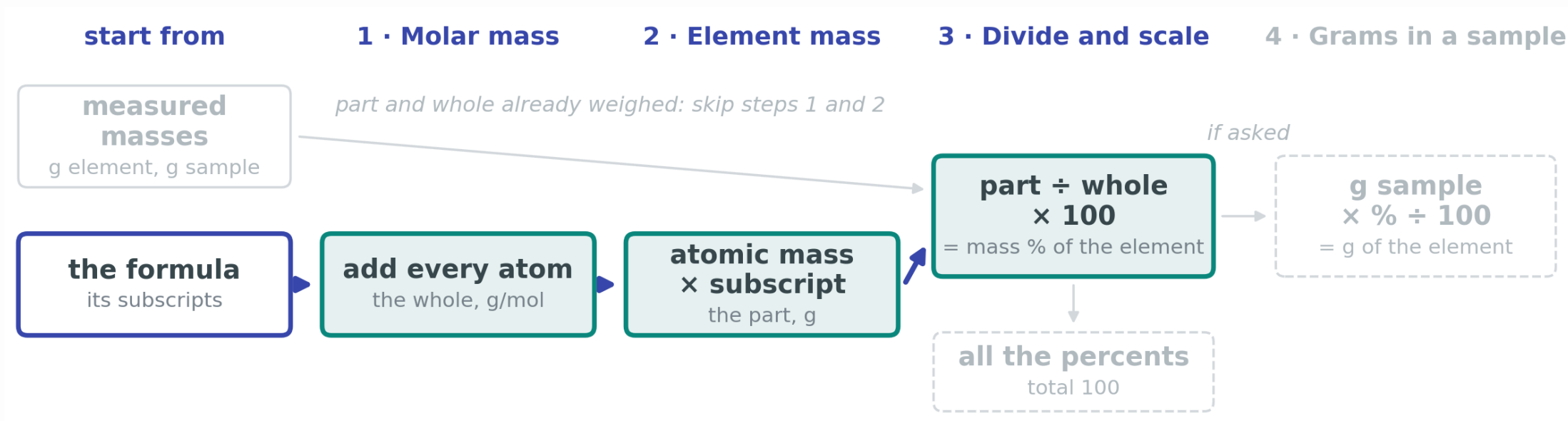
$$\% \text{N} = \frac{28.02 \text{ g N}}{60.06 \text{ g CO}(\text{NH}_2)_2} \times 100 = 46.65\% \text{ N}$$

🧠 **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. ✓

Your turn: sulfur dioxide

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

S 32.07 · O 16.00 g/mol · 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 1

$(\text{NH}_4)_3\text{PO}_4$: 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 1: 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.

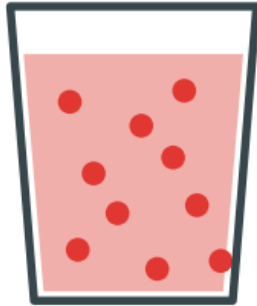
🧠 **Sense check:** The sack holds 2.00 kg of the compound, and 28.19% of that is a bit over a quarter: 564 g. ✓

3 • Molarity

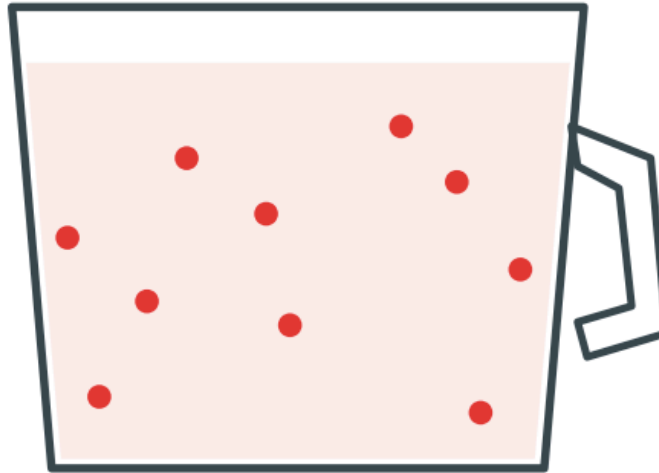
Calculate the molarity of a solution from the amount of solute and the volume of solution, and use it as a conversion factor between solution volume and moles of solute.

Same mix, different strength

one spoonful of mix in each; only the water differs



a small glass
tastes strong

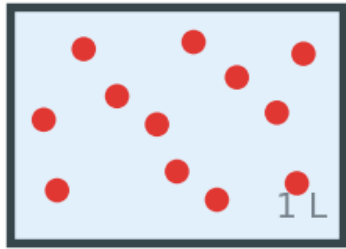


a full pitcher
barely tastes at all

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.

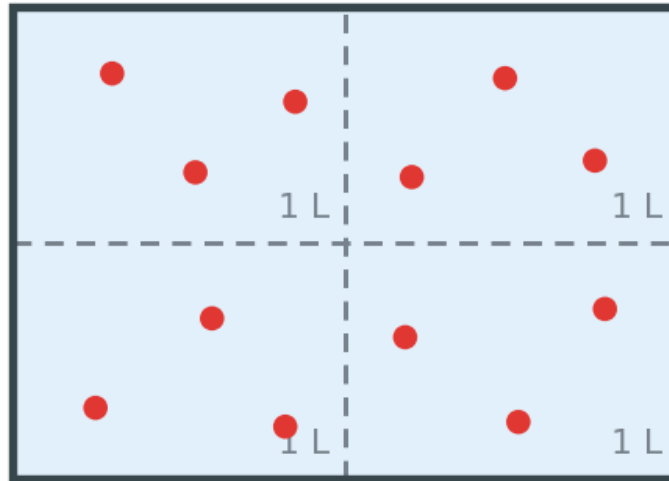
Amount per liter, not total amount

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.

Molarity: moles per liter of solution

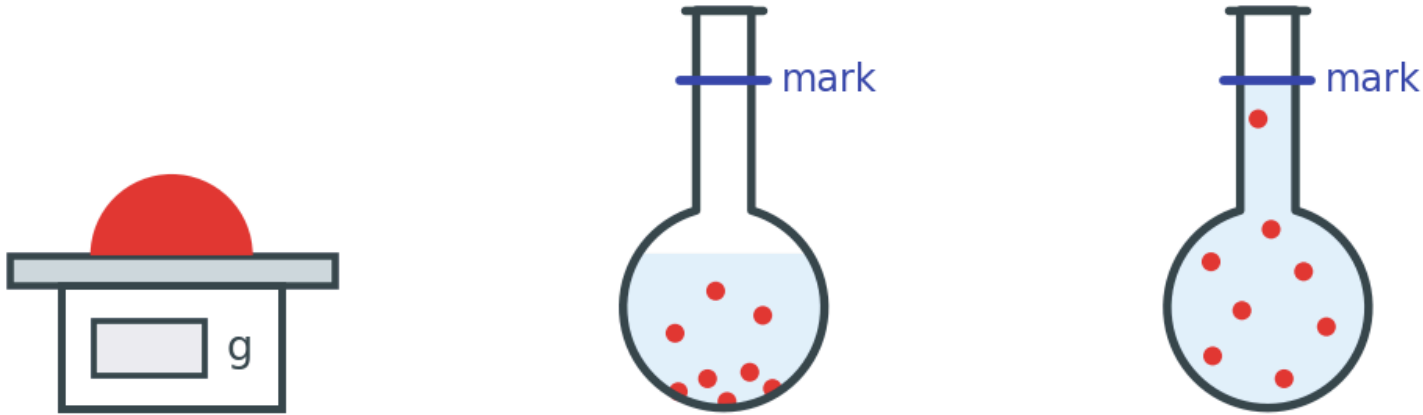
$$M = \text{mol solute} \div \text{L solution}$$

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.



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

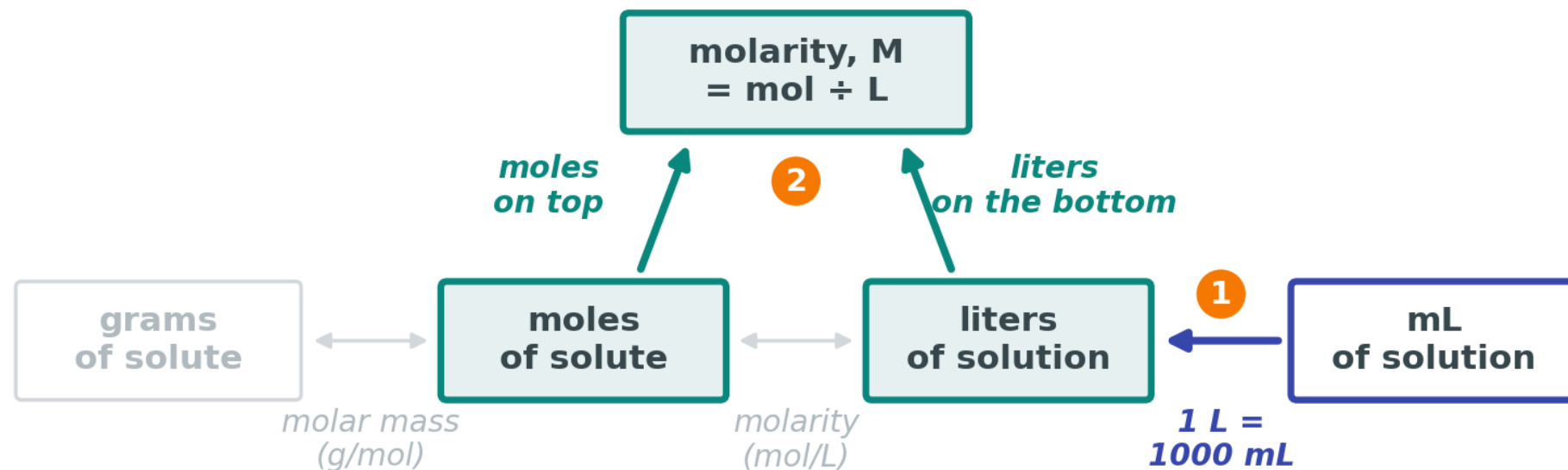
$$M = \frac{0.350 \text{ mol NaCl}}{0.5000 \text{ L soln}} = 0.700 \text{ M NaCl}$$

🧠 **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

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

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

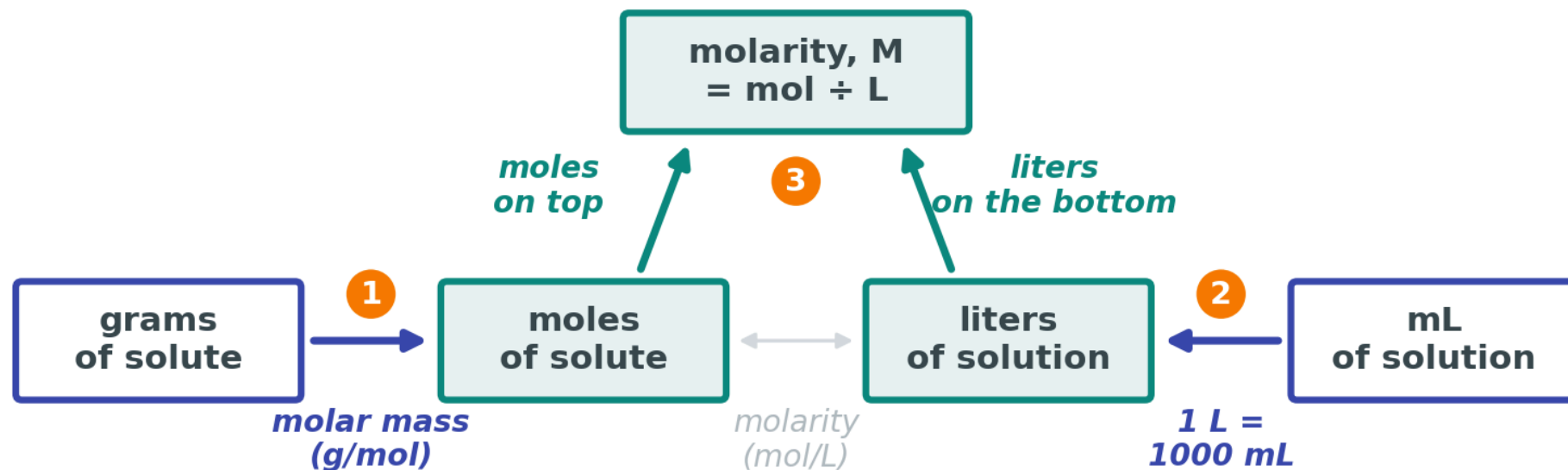
$$9.35 \text{ g KCl} \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} = 0.1254 \text{ mol, then } \frac{0.1254 \text{ mol KCl}}{0.2500 \text{ L soln}} = \mathbf{0.502 \text{ M KCl}}$$

🧠 **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. ✓

Your turn: sodium carbonate

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

given: 21.2 g Na_2CO_3 (105.99 g/mol) · water to the 500.0 mL mark · wanted: M

A wash solution is prepared: **21.2 g of Na_2CO_3** (105.99 g/mol) in a volumetric flask, water to the **500.0 mL** mark.

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

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

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

⚠ 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 1

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

measured: 125.0 mL of solution · 4.46 g KBr left after drying · molar mass KBr 119.00 g/mol

A **125.0 mL** sample of a KBr solution is evaporated to dryness. The dry residue weighs **4.46 g**. What was the molarity of the solution?

- A. 0.300
- B. 0.0375
- C. 0.00468
- D. 0.000300

✓ Practice 1: answer A

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

125.0 mL of solution · 4.46 g KBr (119.00 g/mol) in it · wanted: M

$$4.46 \text{ g KBr} \times \frac{1 \text{ mol KBr}}{119.00 \text{ g KBr}} = 0.0375 \text{ mol, then } \frac{0.0375 \text{ mol KBr}}{0.1250 \text{ L soln}}$$

= 0.300 M KBr (answer A)

B stopped at moles: 0.0375 mol is the amount in the sample, not an amount per liter. C multiplied by the liters: $0.0375 \times 0.1250 = 0.00468$. D divided by milliliters: $0.0375 \div 125.0 = 0.000300$, in mol/mL.

 **Sense check:** An eighth of a liter carries 0.0375 mol, so a full liter carries eight times that: 0.300 mol. 0.300 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: 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 3: 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

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

Write the molarity fraction

Multiply and check

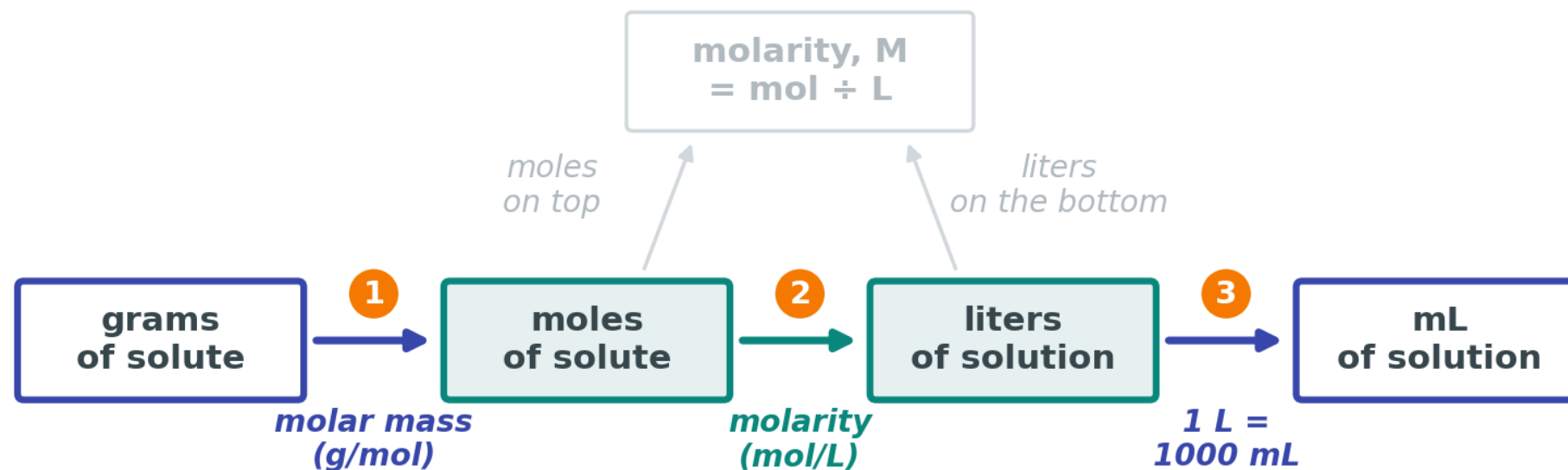
$$15.1 \text{ g KOH} \times \frac{1 \text{ mol KOH}}{56.11 \text{ g KOH}} \times \frac{1 \text{ L soln}}{0.450 \text{ mol KOH}} \times \frac{1000 \text{ mL soln}}{1 \text{ L soln}} = 598 \text{ mL soln}$$

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

✓ Worked example 3: 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:** Three moves: grams to moles, moles to liters, then 0.598 L to 598 mL. ✓



Practice 2

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

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 2: 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

$$10.1 \text{ g KNO}_3 \times \frac{1 \text{ mol KNO}_3}{101.11 \text{ g KNO}_3} \times \frac{1 \text{ L soln}}{0.400 \text{ mol KNO}_3} \times \frac{1000 \text{ mL soln}}{1 \text{ L soln}} = 250. \text{ mL (answer D)}$$

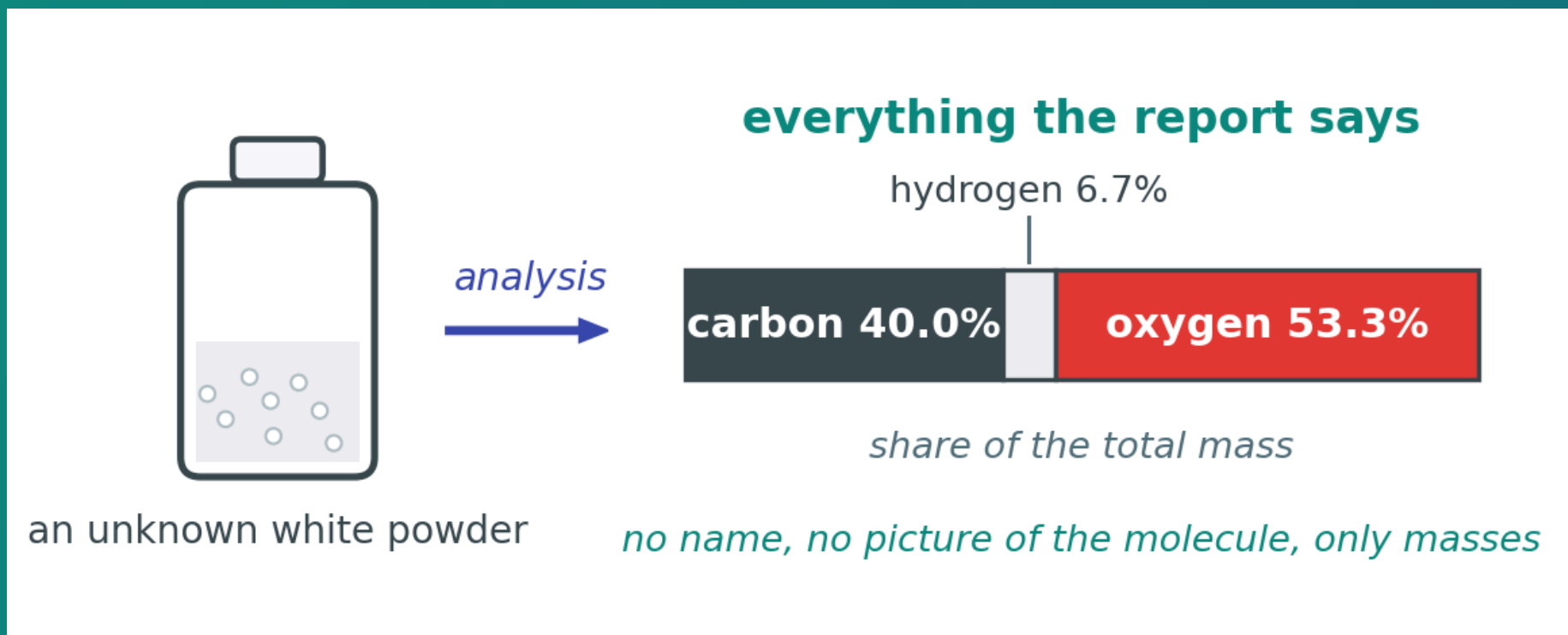
A stopped at liters and relabeled: $10.1 \div 101.11 \div 0.400 = 0.250 \text{ L}$, which is 250. mL, not 0.250 mL. B flipped the molarity fraction: $10.1 \div 101.11 \times 0.400 \times 1000 = 40.0$, and the units, mol²/L, are not a volume. C stopped at moles: $10.1 \div 101.11 = 0.0999 \text{ mol}$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

4 · 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

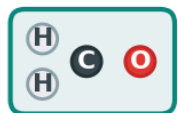


atoms = 6 C : 12 H : 6 O · moles = 6 : 12 : 6, the same ratio · grams = 72.1 : 12.1 : 96.0, not the same

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

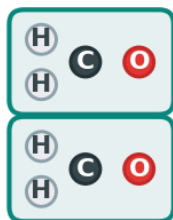
one repeating unit, 1 C : 2 H : 1 O



x 1



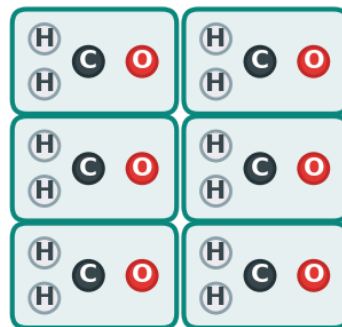
formaldehyde



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.

The 100 g sample

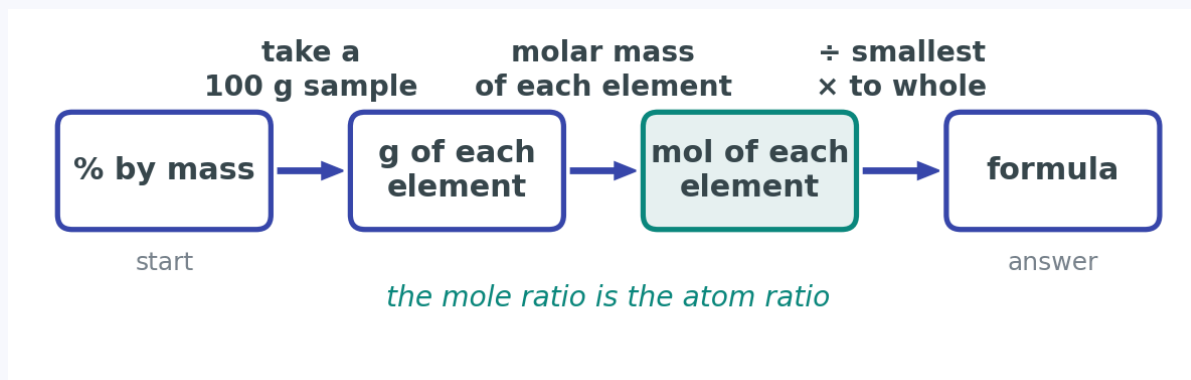
40.0% C $\rightarrow$ 40.0 g C in every 100 g of compound

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



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

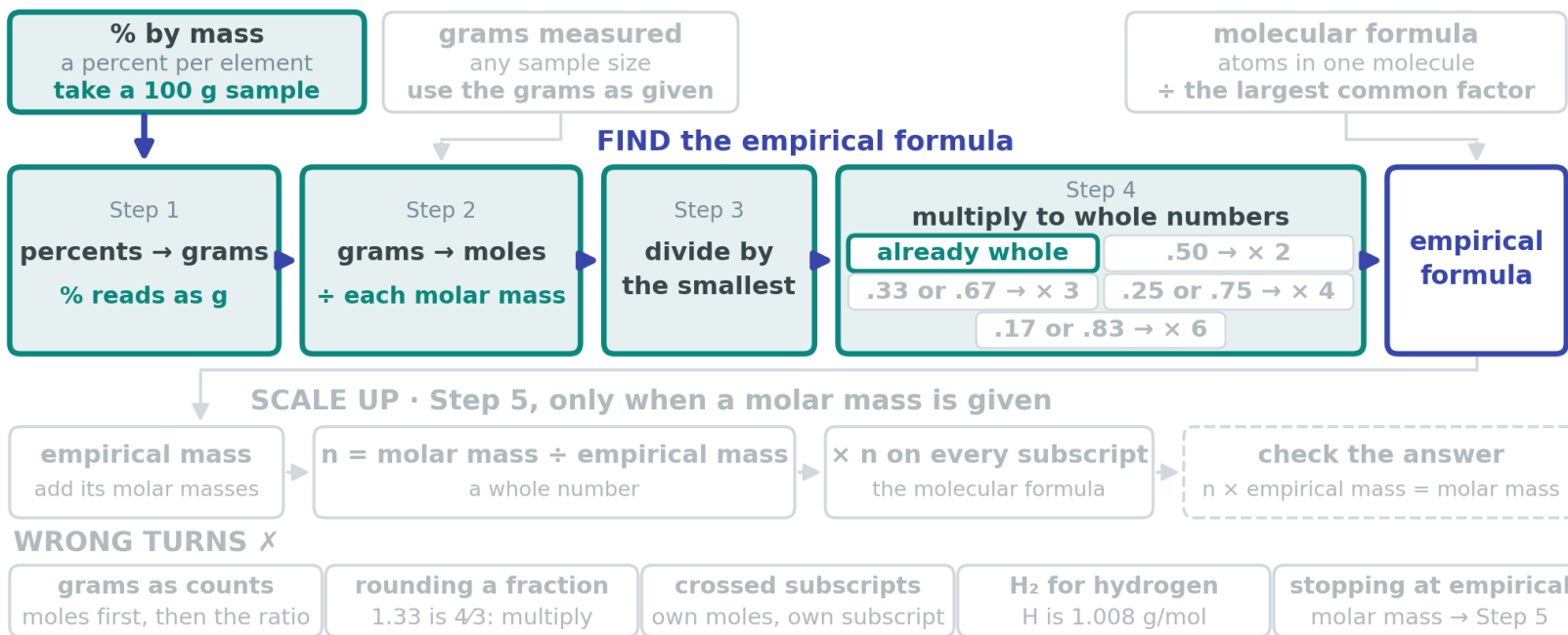
1 C : 2 H : 1 O · the smallest whole-number atom ratio

🧠 **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:** Steps 1 to 4 from the percents: 1 : 2 : 1 came out already whole, so Step 4 had nothing to clear. ✓



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

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

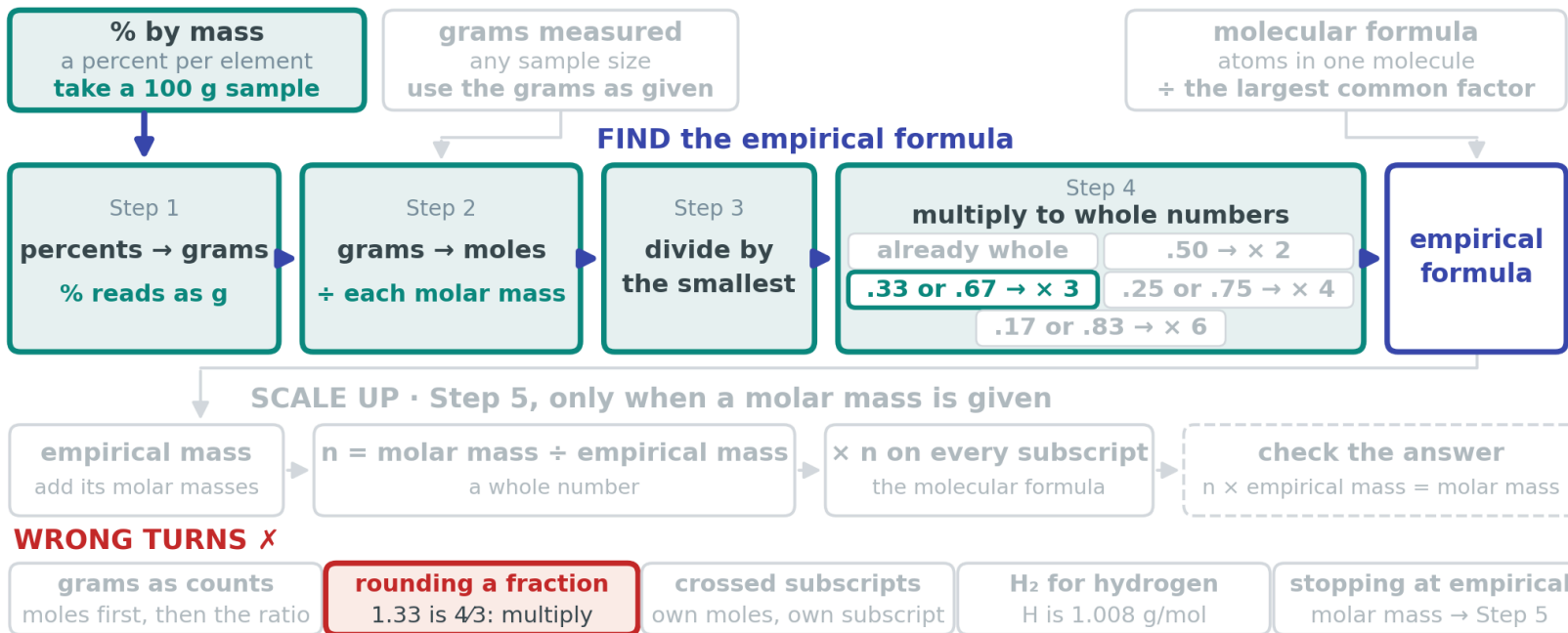
 **Sense check:** Fe₃O₄ rebuilds to 72.4% Fe: the data agree. ✓

✓ 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 $1 : 1 \rightarrow \text{FeO}$, 77.7% Fe against 72.4% measured ✕

$.50 \rightarrow \times 2 \cdot .33 \text{ or } .67 \rightarrow \times 3 \cdot .25 \text{ or } .75 \rightarrow \times 4$
multiply every element's count by the same factor · 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 1

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 1: 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.

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

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

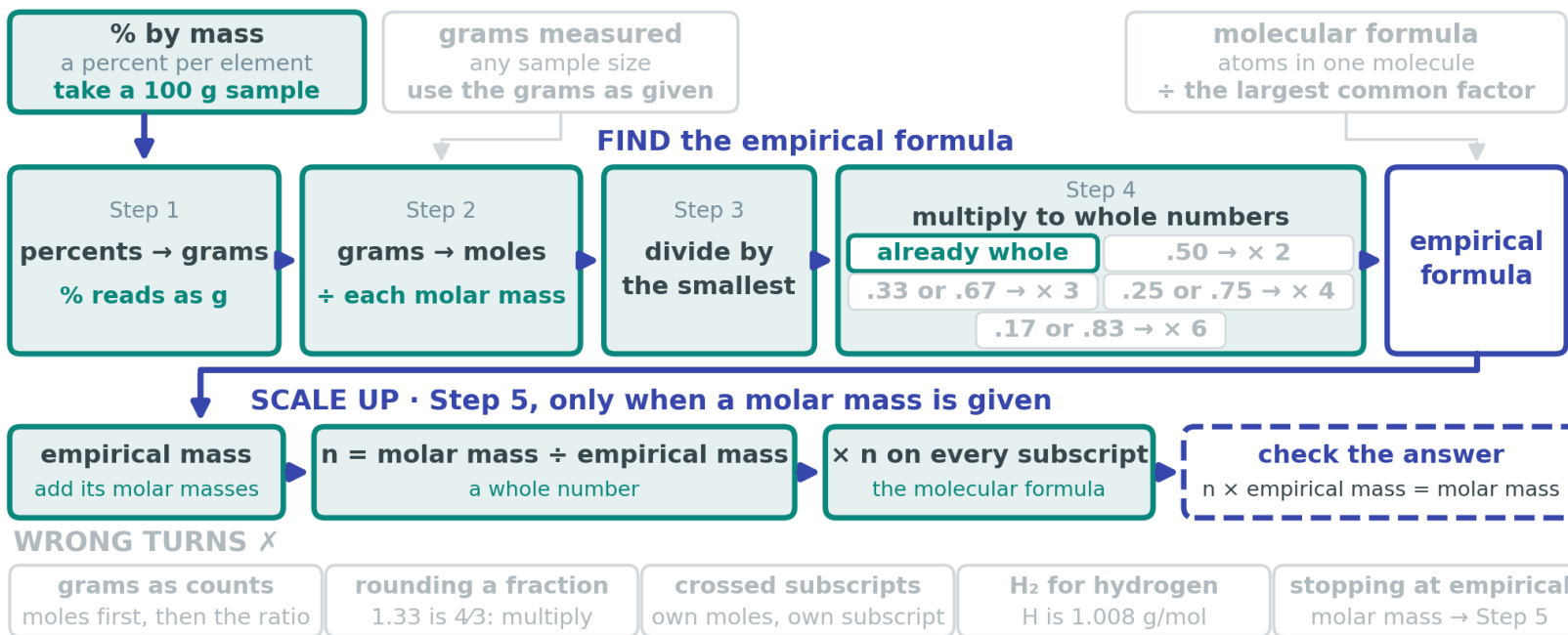
 **Sense check:** n landed whole, and $6 \times 30.03 = 180.2$ rebuilds the molar mass. An unknown white powder with these numbers is glucose, blood sugar. ✓

✓ 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 2

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 2: 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.

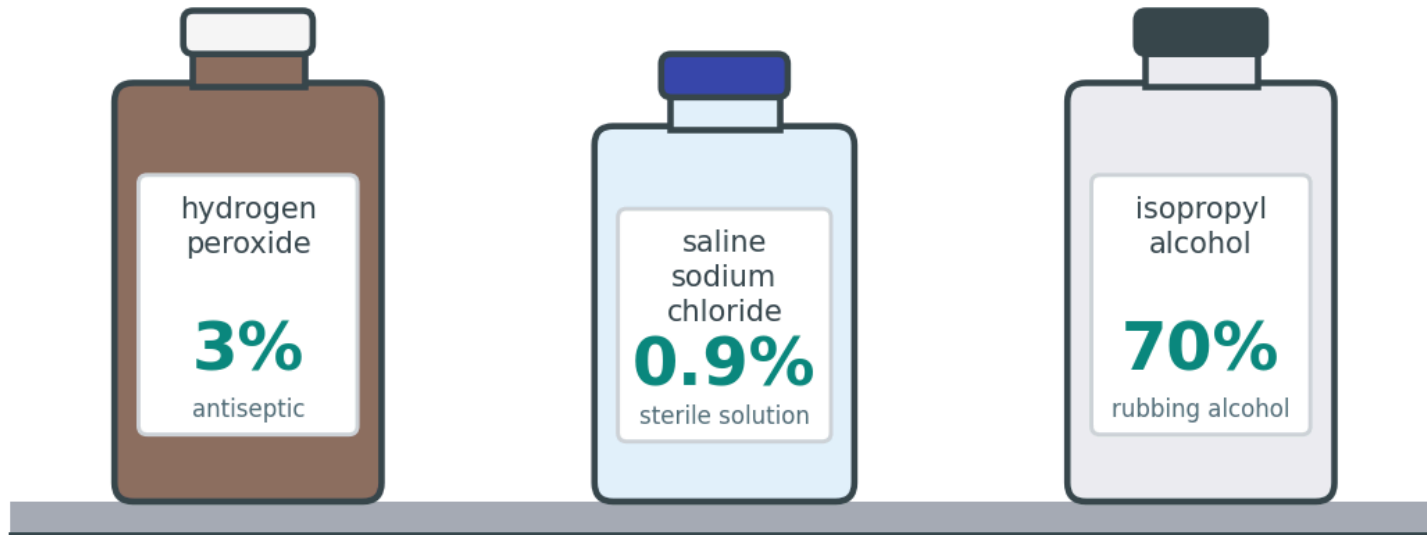
🧠 **Sense check:** n = 6 lands whole: 6 × 13.02 = 78.1. B (75.08) and D (84.16) sit close: mass alone cannot decide. ✓

5 • 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

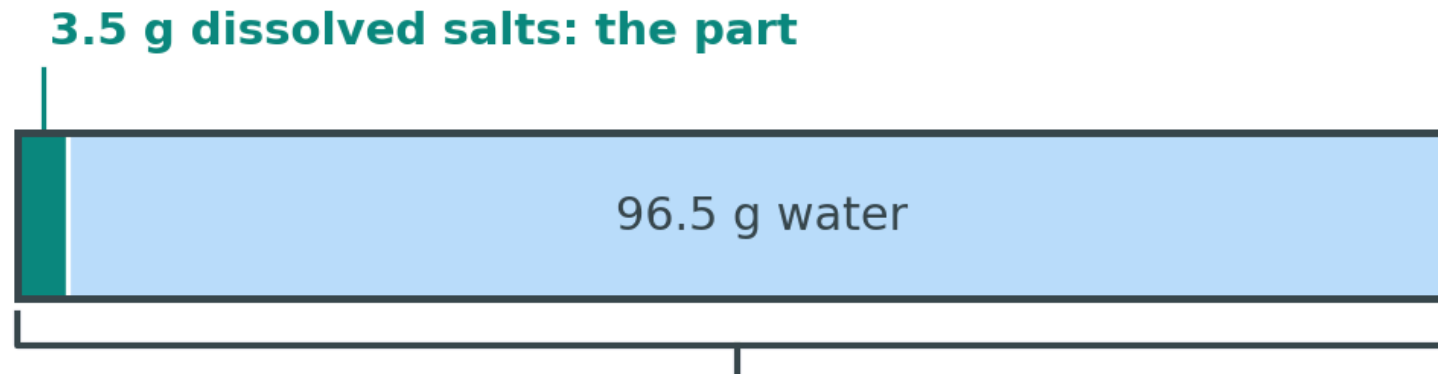
one drugstore shelf, three labels, three percents



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.

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.**



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

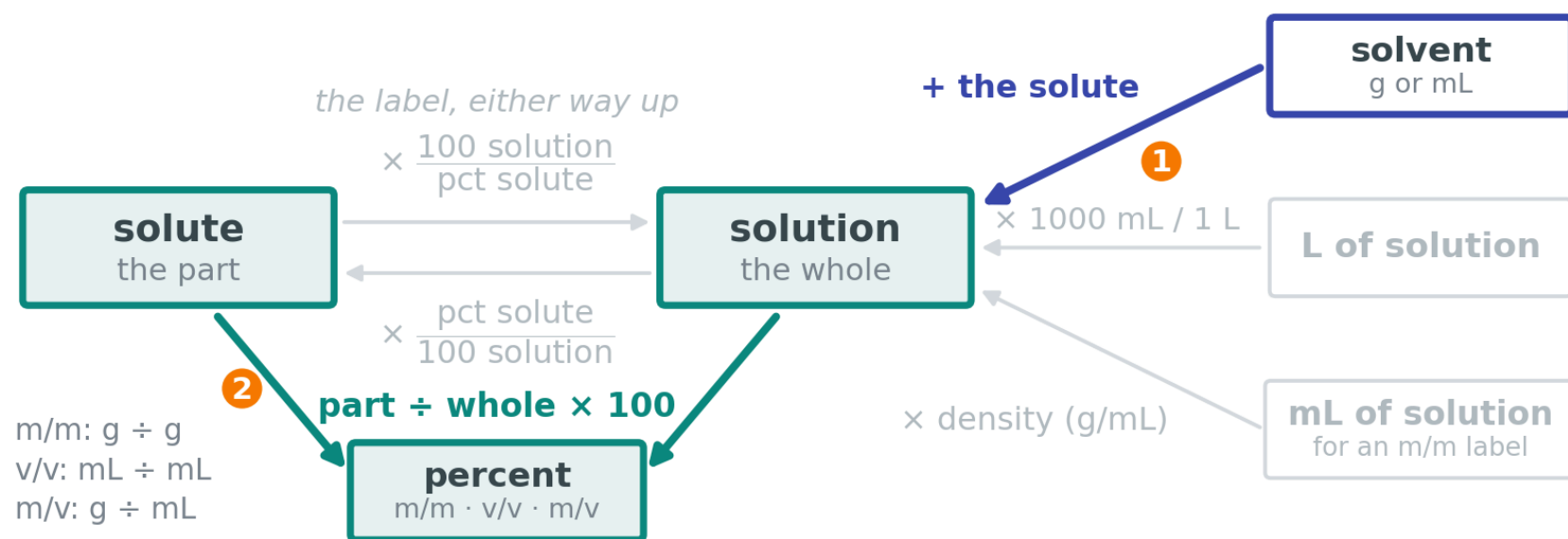
$$\frac{25.0 \text{ g sucrose}}{125.0 \text{ g solution}} \times 100 = \mathbf{20.0\% (m/m)}$$

 **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 $\times 100$. ✓



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

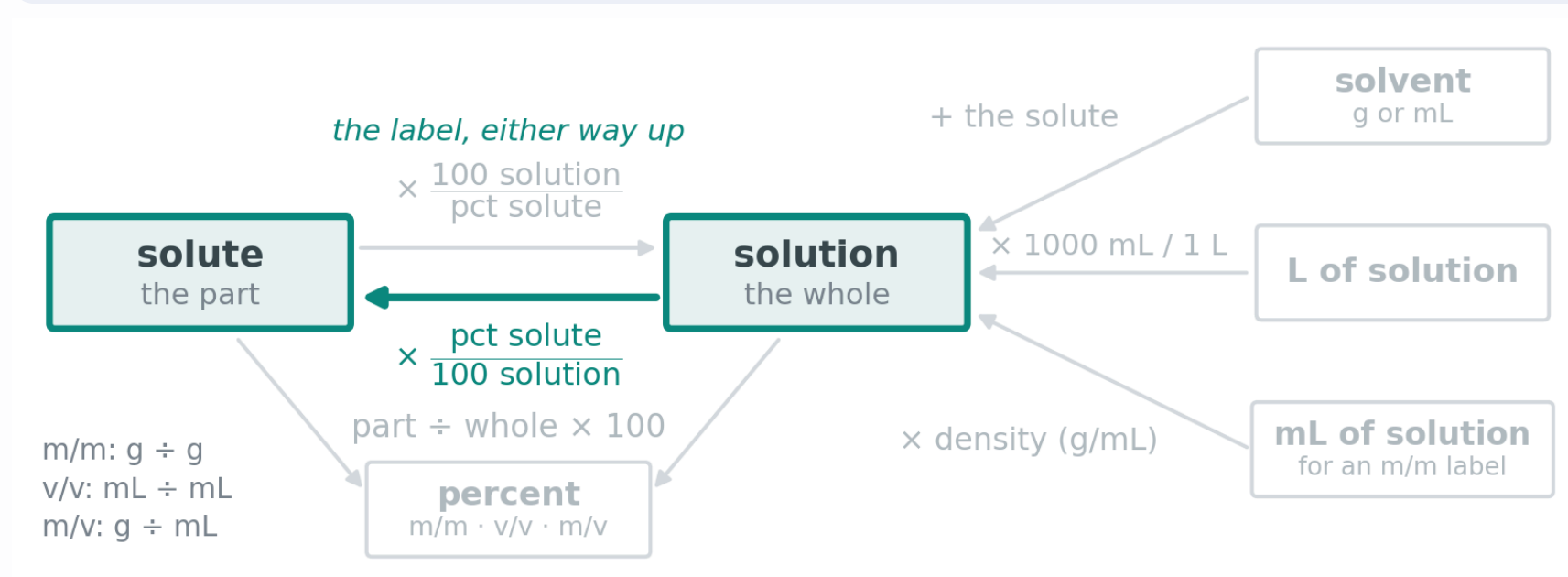
Step 4 · Multiply and check

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

 **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 g H₂O₂ = 100 g solution
 given: 250.0 g solution · found: 7.5 g H₂O₂



🧠 **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 1

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 1: 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

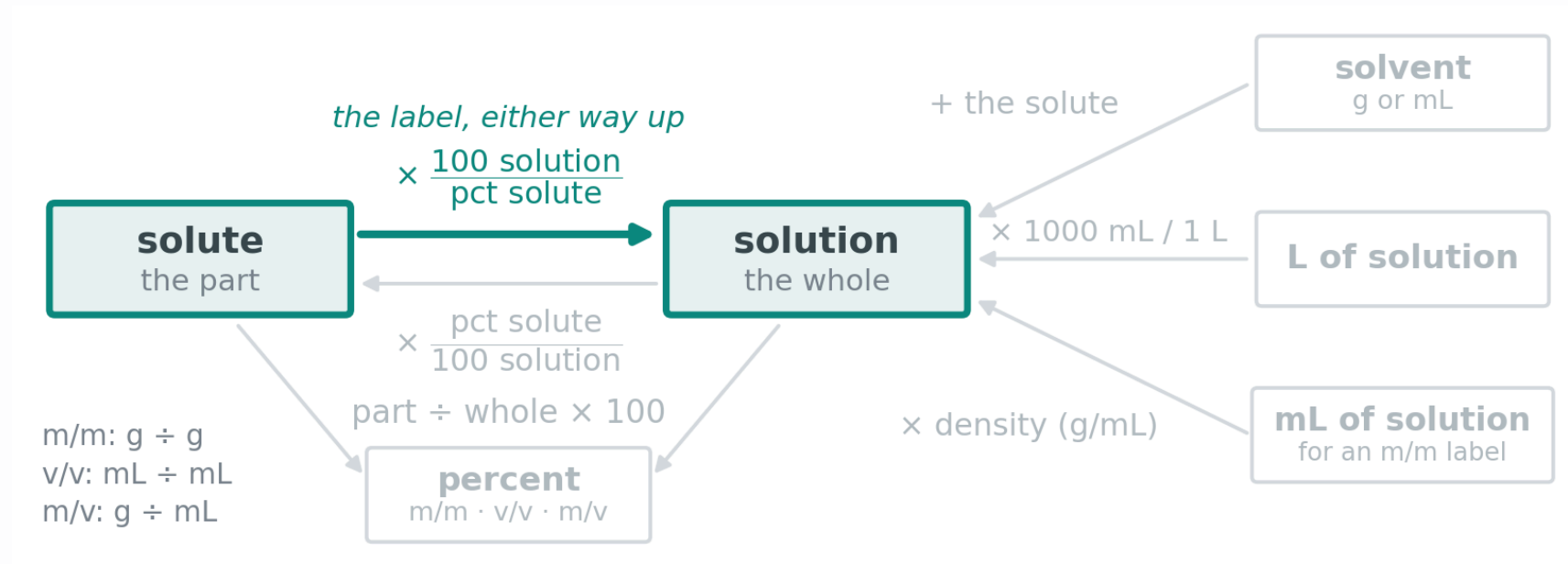
Step 4 · Multiply and check

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

 **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 2

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 2: answer B

28.0% (m/m): 28.0 g NH₃ = 100 g solution · 0.900 g solution = 1 mL solution

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

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

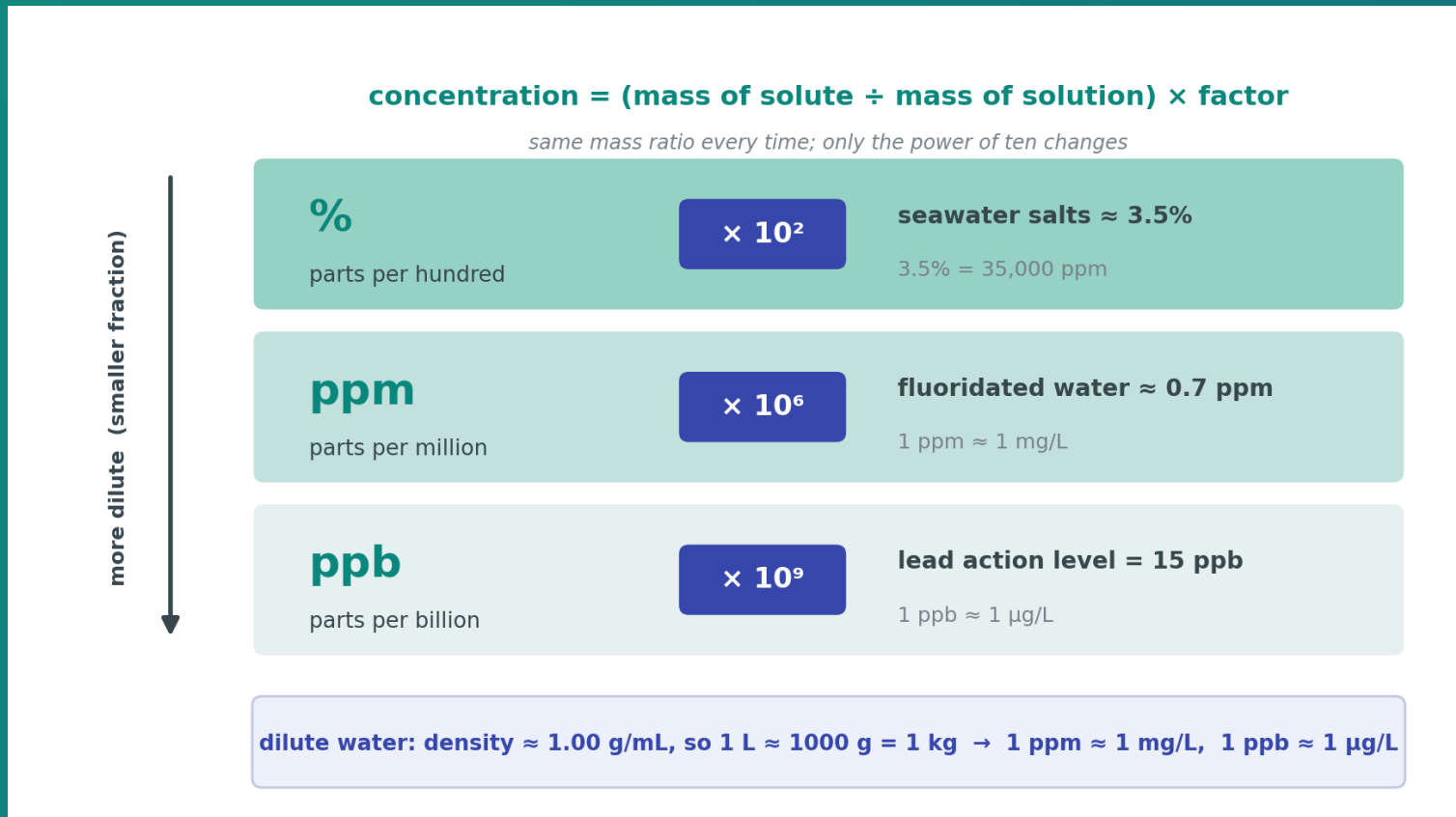
A skipped the density: $250. \times 28.0 / 100 = 70.0$ treats mL of solution as grams. C stopped halfway: $250. \times 0.900 = 225 \text{ g}$ is the whole solution, not the NH₃ in it. D flipped the density: $250. \div 0.900 = 278 \text{ g}$, then $\times 28.0 / 100 = 77.8$.

 **Sense check:** The liquid is lighter than water, so 250. mL weighs less than 250 g, and the NH₃ lands below the 70.0 g a density of 1 would give: 63.0 g. ✓

6 · Parts per Million & Parts per Billion

Calculate a solution's concentration in ppm or ppb from the masses of solute and solution, use the dilute-water shortcut $1 \text{ ppm} \approx 1 \text{ mg/L}$ ($1 \text{ ppb} \approx 1 \text{ }\mu\text{g/L}$), and use a ppm value as a conversion factor between volume of water and mass of solute.

When a percent is far too big



A water report lists lead at 15 ppb and fluoride at 0.7 ppm, never as a percent. A solute this dilute needs a denominator bigger than a hundred.

Below one part in a hundred

A percent is parts per **hundred**. A trace solute never reaches one part in a hundred, so the same mass ratio is scaled to a **larger** whole.

ppm = parts per million (10^6) · ppb = parts per billion (10^9)

both are mass ratios: mass of solute over mass of solution: the same fraction a percent uses, times a bigger power of ten

The two formulas

Each is the mass fraction times its power of ten. The denominator is the **whole solution**, solute plus solvent.

$$\text{ppm} = (\text{mass solute} / \text{mass solution}) \times 10^6$$
$$\text{ppb} = (\text{mass solute} / \text{mass solution}) \times 10^9, \text{ a thousand times finer}$$

The mass units cancel, so ppm and ppb carry no unit of their own.

The water shortcut: 1 ppm $\approx$ 1 mg/L

A **dilute aqueous** solution has density near 1.00 g/mL, so 1 L weighs about 1000 g.

$$1 \text{ ppm} \approx 1 \text{ mg/L} \cdot 1 \text{ ppb} \approx 1 \text{ }\mu\text{g/L}$$

1 mg in 1 kg = 1 mg per 1000 g = 1 part in 10^6 = 1 ppm; the density $\approx$ 1.00 g/mL is what makes L $\leftrightarrow$ kg work

A ppm then reads as mg/L. A dense brine or a nonaqueous solvent breaks the shortcut.



The method

1. **Mass fraction, same units:** solute mass over solution mass.
2. **Multiply by the power of ten:** 10^6 for ppm, 10^9 for ppb.
3. **In dilute water:** 1 ppm reads as 1 mg/L; $\text{mg/L} \times \text{liters}$ gives the mass.



Worked example: ppm from masses

0.0030 g copper dissolved in 250. g of water sample

given: 0.0030 g solute · 250. g solution · wanted: ppm

Old copper plumbing leaches into a **250. g** water sample, adding **0.0030 g of copper**.
What is the copper concentration in ppm?

Both masses are already in grams: set up the mass fraction, then scale.

✓ Worked example: solution

Step 1 • Mass fraction, same units

$$\frac{0.0030 \text{ g Cu}}{250. \text{ g solution}} = 1.2 \times 10^{-5}$$

Step 2 • Multiply by the power of ten

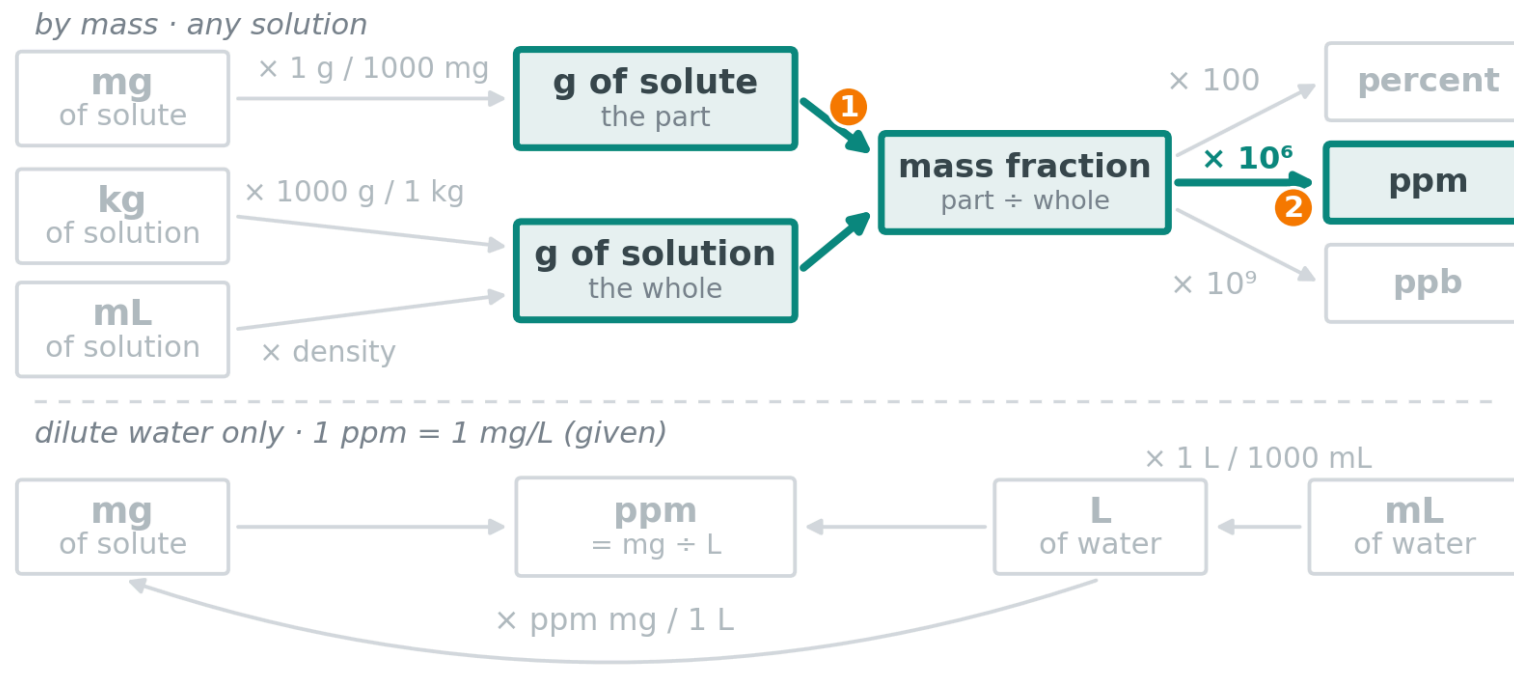
$$1.2 \times 10^{-5} \times 10^6 = 12 \text{ ppm}$$
$$0.0030 / 250 \times 10^6 = 12 \text{ ppm of copper}$$

🧠 **Sense check:** 12 ppm $\approx$ 12 mg/L: 12 mg of copper in each liter of this water. If you had used 10^9 you would report 12,000 (that is ppb), and 10^3 would give 0.012 (per-thousand). The $\times 10^6$ is what makes it ppm. ✓

✓ Worked example: the route on the map

0.0030 g copper dissolved in 250. g of water sample

given: 0.0030 g solute · 250. g solution · found: 12 ppm



🧠 **Sense check:** Two moves: part over whole, then $\times 10^6$. No unit change was needed. ✓

Your turn: ppm as a conversion factor

fluoridated tap water: 0.70 ppm fluoride $\approx$ 0.70 mg/L

given: 0.70 ppm $\cdot$ 2.0 L of water $\cdot$ wanted: mg of fluoride

A city fluoridates its water to **0.70 ppm**. How many milligrams of fluoride are in a **2.0 L** pitcher? Read ppm as mg/L, then multiply by the volume.

$$2.0 \text{ L} \times \frac{\text{mg F}^-}{1 \text{ L}} = \text{mg F}^-$$

$$2.0 \text{ L} \times \frac{0.70 \text{ mg F}^-}{1 \text{ L}} = 1.4 \text{ mg F}^-$$

How small is one ppm?

1 ppm = 1 mg in 1 kg = 1 second in ~11.6 days

1 ppb is a thousand times smaller still: about 1 second in 32 years

Use a **percent** for everyday concentrations (parts per hundred). Reach for **ppm** when the solute is a trace (dissolved minerals, fluoride, a pollutant) and **ppb** for the barely-there, like lead limits in drinking water.

⚠ Where this goes wrong

Wrong power of ten. $0.0030 / 250 \times 10^9 = 12,000$ is ppb, not ppm; $\times 10^3 = 0.012$ is per-thousand. ppm uses $\times 10^6 \rightarrow 12$.

Dividing by the solvent alone. The denominator is the whole solution, solute + solvent, the same whole a percent uses.

Forgetting the volume. 0.70 ppm is 0.70 mg per liter; a 2.0 L pitcher holds $0.70 \times 2.0 = 1.4$ mg, not 0.70 mg.

Using the shortcut on a non-dilute or nonaqueous solution. 1 ppm $\approx$ 1 mg/L needs density $\approx$ 1.00 g/mL; a heavy brine or an organic solvent breaks $1 \text{ L} \approx 1 \text{ kg}$.

 Practice

A **160.0 mL** sample of fruit-juice concentrate contains **4.16 mg** of dissolved tin. The concentrate has a density of **1.30 g/mL**. What is the tin concentration in ppm?

- A. 2.00×10^{-5}
- B. 20.0
- C. 26.0
- D. 33.8

✓ Practice: answer B

$$160.0 \text{ mL} \times 1.30 \text{ g/mL} = 208.0 \text{ g solution} \cdot 4.16 \text{ mg} = 0.00416 \text{ g tin}$$

1 ppm $\approx$ 1 mg/L needs a density near 1.00 g/mL; this concentrate is 1.30 g/mL

$$\frac{0.00416 \text{ g tin}}{208.0 \text{ g solution}} \times 10^6 = \mathbf{20.0 \text{ ppm (answer B)}}$$

A stopped at the mass fraction: $0.00416 \div 208.0 = 2.00 \times 10^{-5}$, never scaled by 10^6 . C used the water shortcut: $4.16 \text{ mg} \div 0.1600 \text{ L} = 26.0 \text{ mg/L}$, read as ppm; that needs $1 \text{ L} \approx 1 \text{ kg}$. D flipped the density: $160.0 \div 1.30 = 123 \text{ g}$, and $0.00416 \div 123 \times 10^6 = 33.8$.

 **Sense check:** A dense liquid packs more grams into each mL, so the same tin sits in more solution mass: 20.0 ppm, below the shortcut's 26.0. ✓



Check yourself

1. A 500. g water sample holds 0.010 g of iron. What is the iron concentration in ppm?
2. A pond is treated to 2.5 ppm of a chemical. How many milligrams are in 4.0 L of pond water? State the shortcut you used.

The same mass fraction runs through percent, ppm, and ppb: only the power of ten shifts with how dilute the solute is. Next, these mole tools drive reactions: a balanced equation is a recipe written in moles, and stoichiometry reads that recipe.

Can you...?

- ☐ calculate the molar mass of any compound from its formula?
- ☐ convert among grams, moles, and particles using molar mass and Avogadro's number?
- ☐ calculate the percent composition of a compound from its formula?
- ☐ calculate molarity and use it as a conversion factor between solution volume and moles of solute?
- ☐ determine empirical and molecular formulas from experimental data?
- ☐ express solution concentration as mass, volume, or mass-volume percent and use it as a conversion factor?

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