

Solutions

Preparation for General Chemistry · Dr. Karmach

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

- Identify solute and solvent, write dissolving equations, and predict solubility with like dissolves like
- Explain the dissolving process through the attractions broken and formed, and why dissolving can cool or warm the solution
- Classify a solution as unsaturated, saturated, or supersaturated and predict how temperature and pressure (Henry's law) change solubility
- Calculate and use molarity, percent concentration, ppm/ppb, and molality, and dilute a solution with $M_1V_1 = M_2V_2$
- Carry a solution's volume and molarity through reaction stoichiometry
- Describe the colligative properties and calculate freezing-point depression and boiling-point elevation with $\Delta T = i \cdot K \cdot m$

Today's route

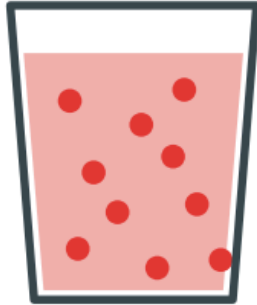
1. **Molarity**
2. **Dilution**
3. **Like Dissolves Like**
4. **Saturation & Solubility Curves**
5. **Henry's Law**
6. **Percent Concentration**
7. **Molality & Mole Fraction**
8. **Freezing Point & Boiling Point**
9. **Parts per Million & Parts per Billion**
10. **Solution Stoichiometry**

1 • Molarity

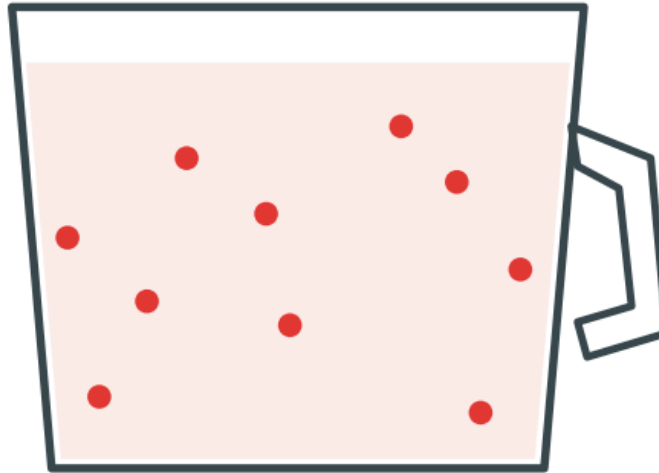
Calculate the molarity of a solution from the amount of solute and the volume of solution, and use it as a conversion factor between solution volume and moles of solute, including the grams to weigh out to prepare a solution.

Same mix, different strength

one spoonful of mix in each; only the water differs



a small glass
tastes strong

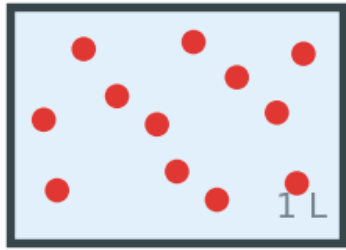


a full pitcher
barely tastes at all

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.

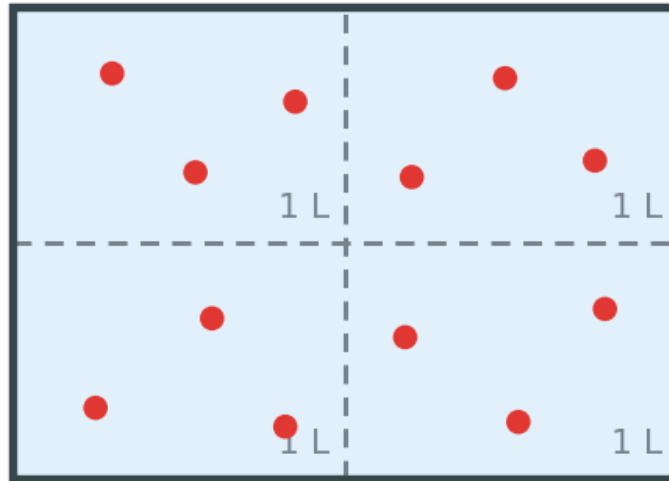
Amount per liter, not total amount

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.

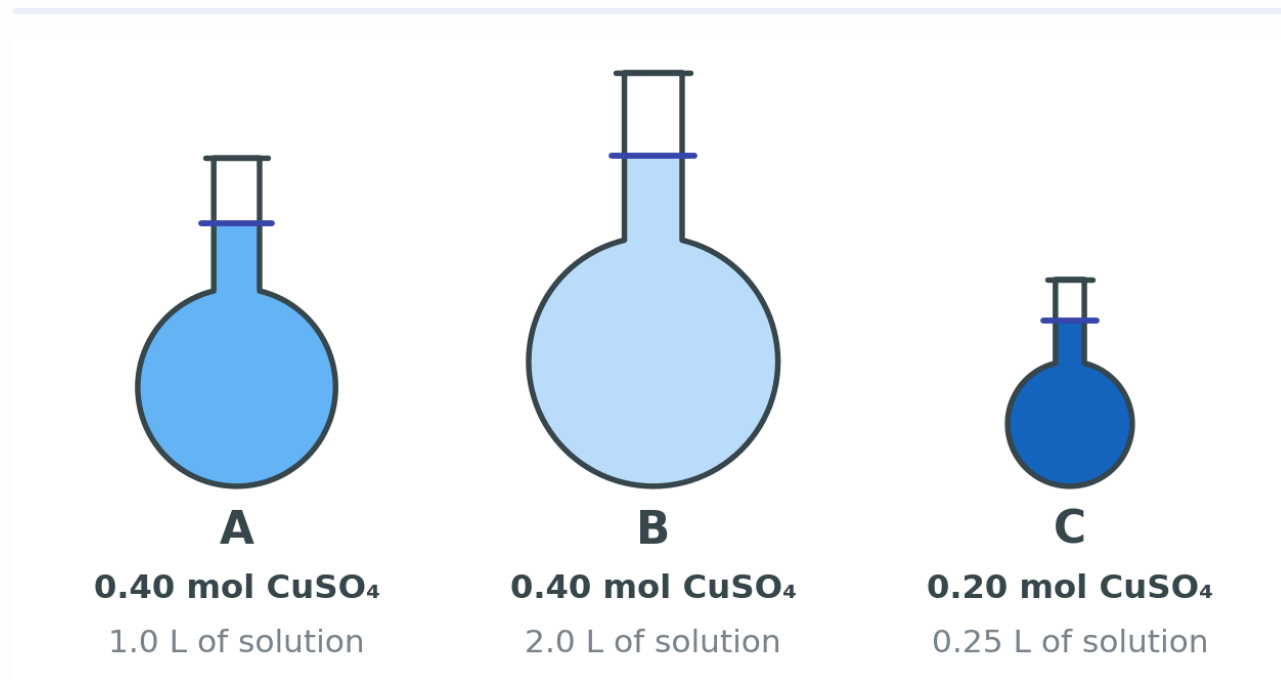
Solute, solvent, solution

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

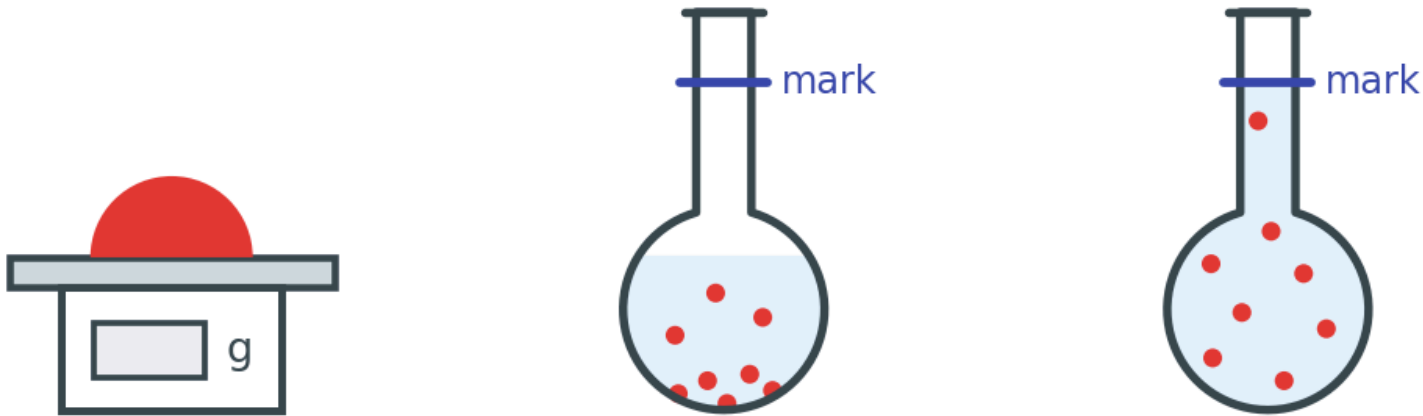
$$M = \text{mol solute} \div \text{L solution}$$

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

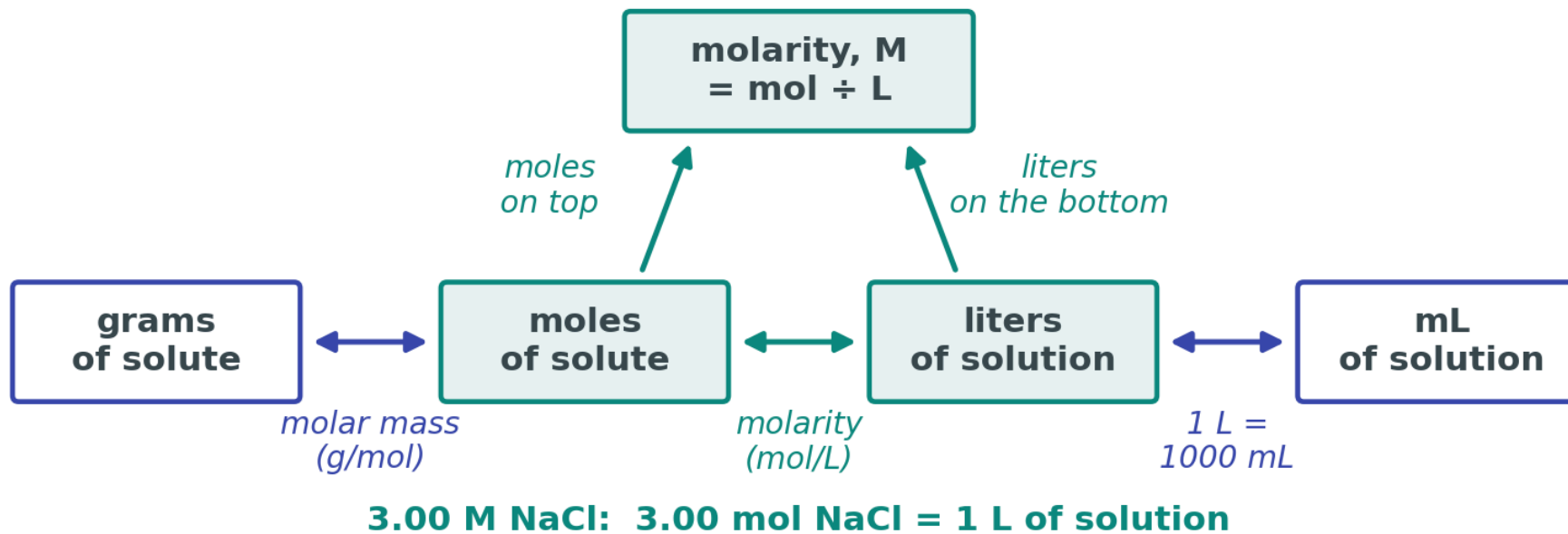
The solute takes up room. Prepare the solution in a volumetric flask: solute in first, then water to the mark. The mark reads the volume of finished solution, solute included.



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

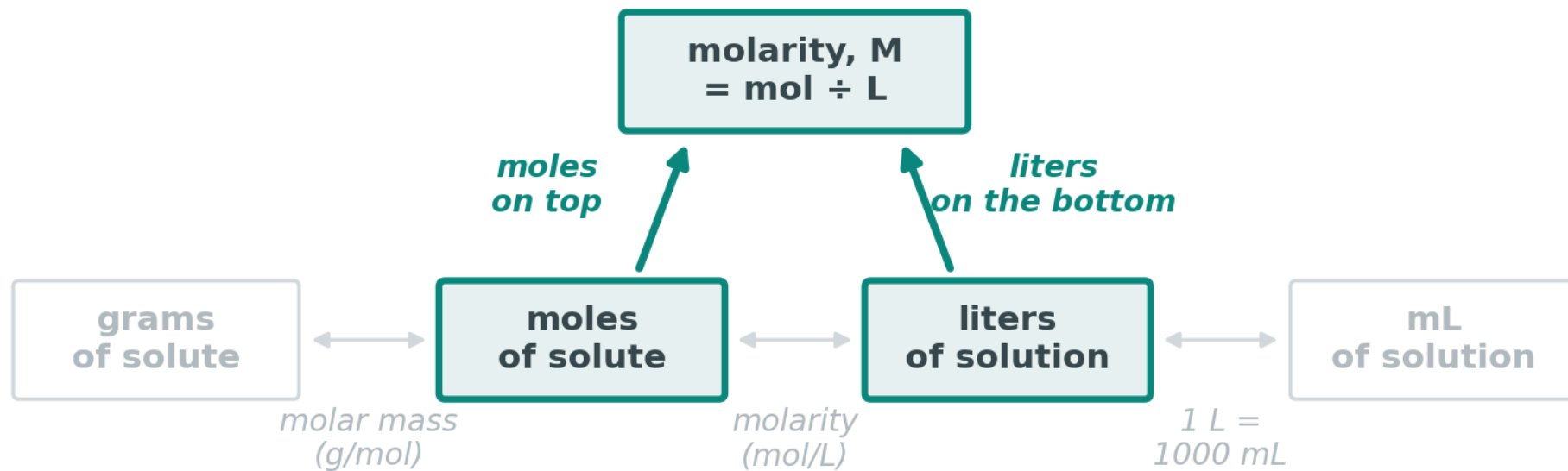
$$M = \frac{0.300 \text{ mol glucose}}{1.50 \text{ L soln}} = 0.200 \text{ M glucose}$$

🧠 **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

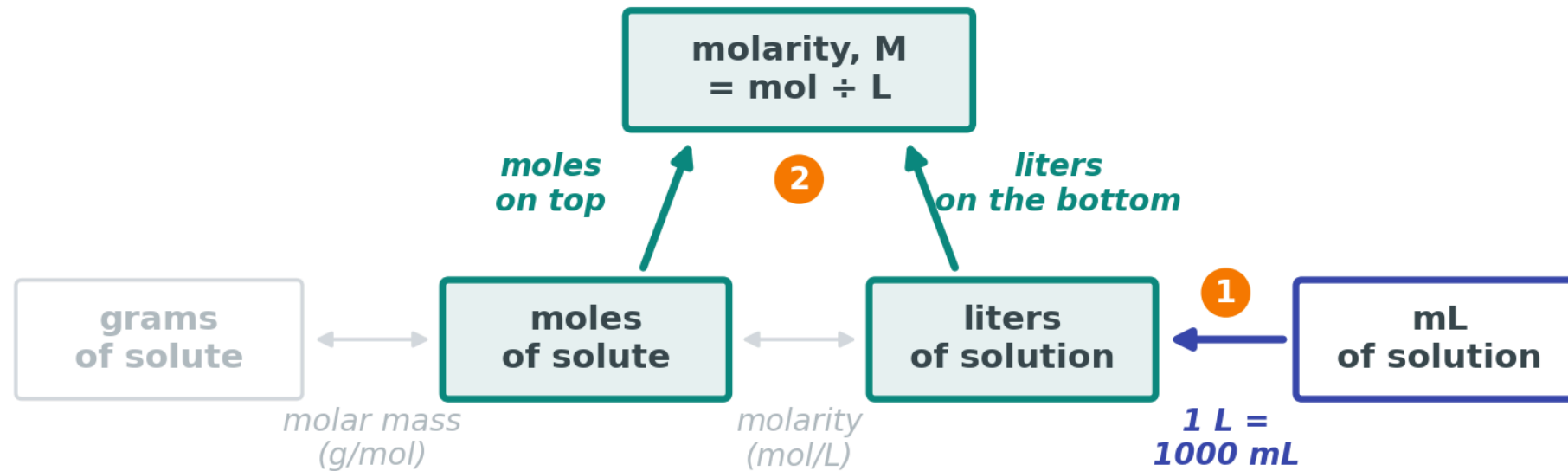
$$M = \frac{0.350 \text{ mol NaCl}}{0.5000 \text{ L soln}} = 0.700 \text{ M NaCl}$$

🧠 **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

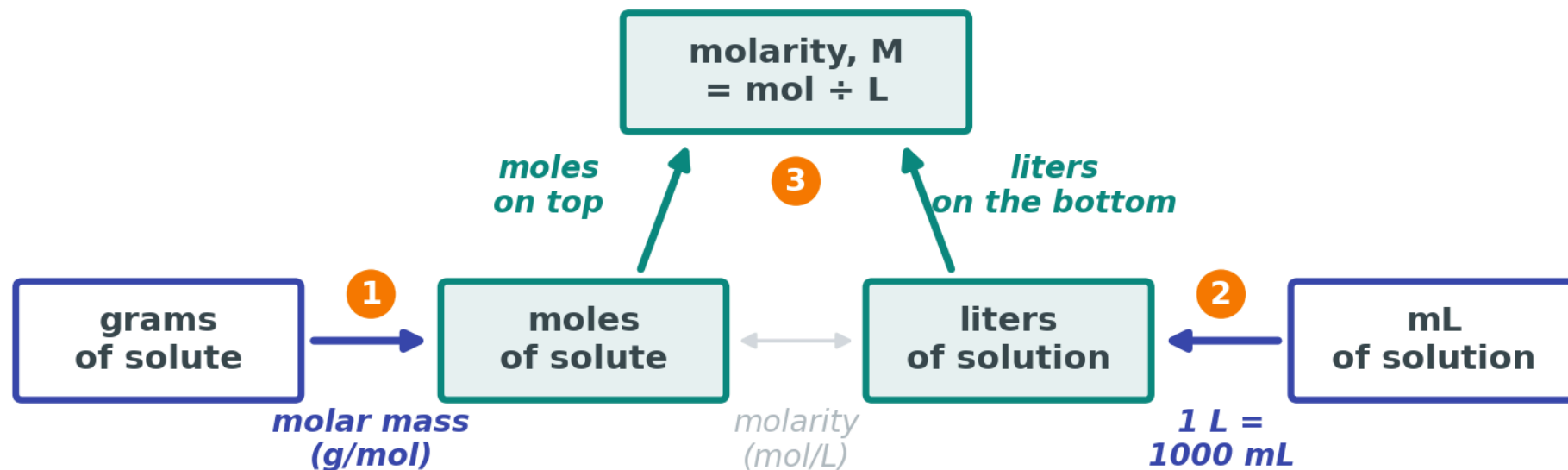
$$9.35 \text{ g KCl} \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} = 0.1254 \text{ mol, then } \frac{0.1254 \text{ mol KCl}}{0.2500 \text{ L soln}} = \mathbf{0.502 \text{ M KCl}}$$

🧠 **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 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.

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

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

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

⚠ 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 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

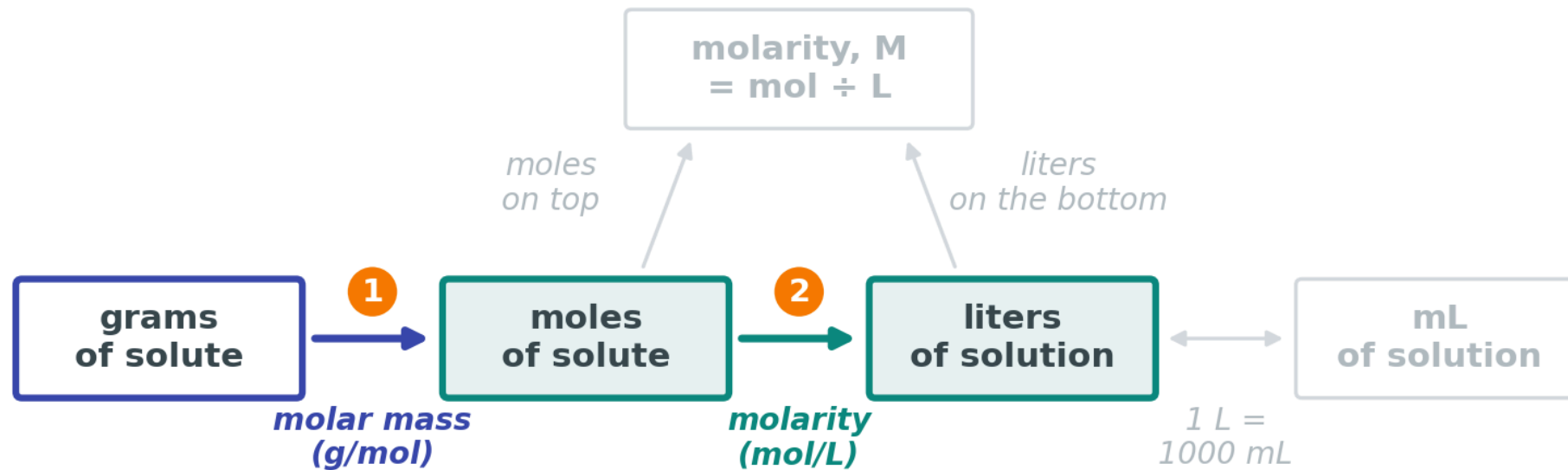
$$351 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \times \frac{1 \text{ L soln}}{3.00 \text{ mol NaCl}} = 2.00 \text{ L soln}$$

🧠 **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

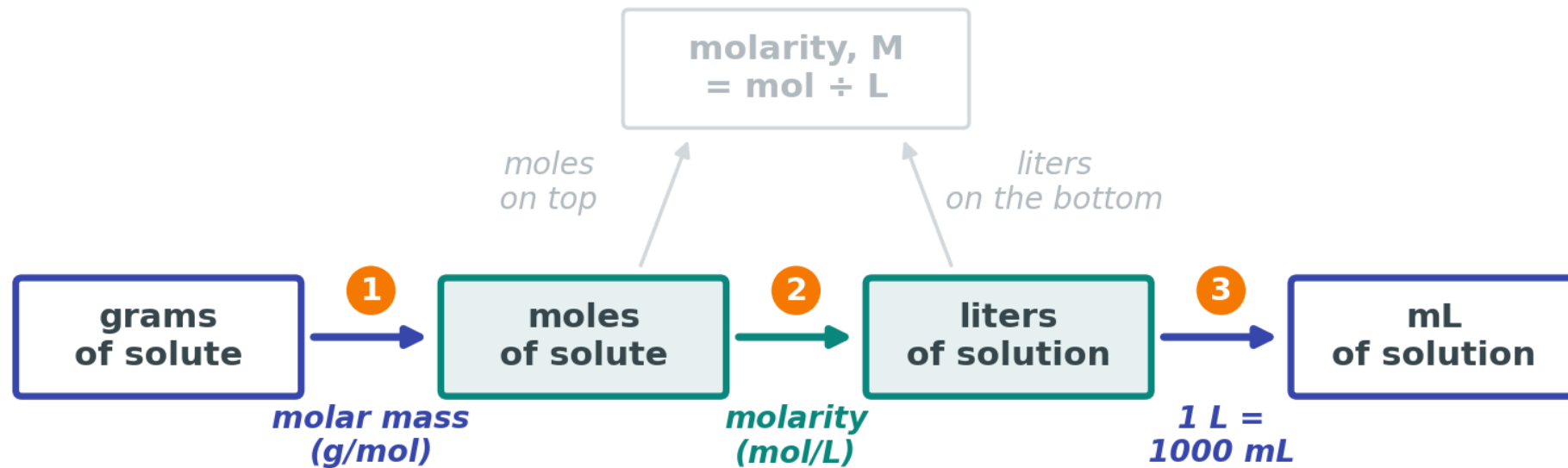
$$15.1 \text{ g KOH} \times \frac{1 \text{ mol KOH}}{56.11 \text{ g KOH}} \times \frac{1 \text{ L soln}}{0.450 \text{ mol KOH}} \times \frac{1000 \text{ mL soln}}{1 \text{ L soln}} = 598 \text{ mL soln}$$

 **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

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

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 ✓



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

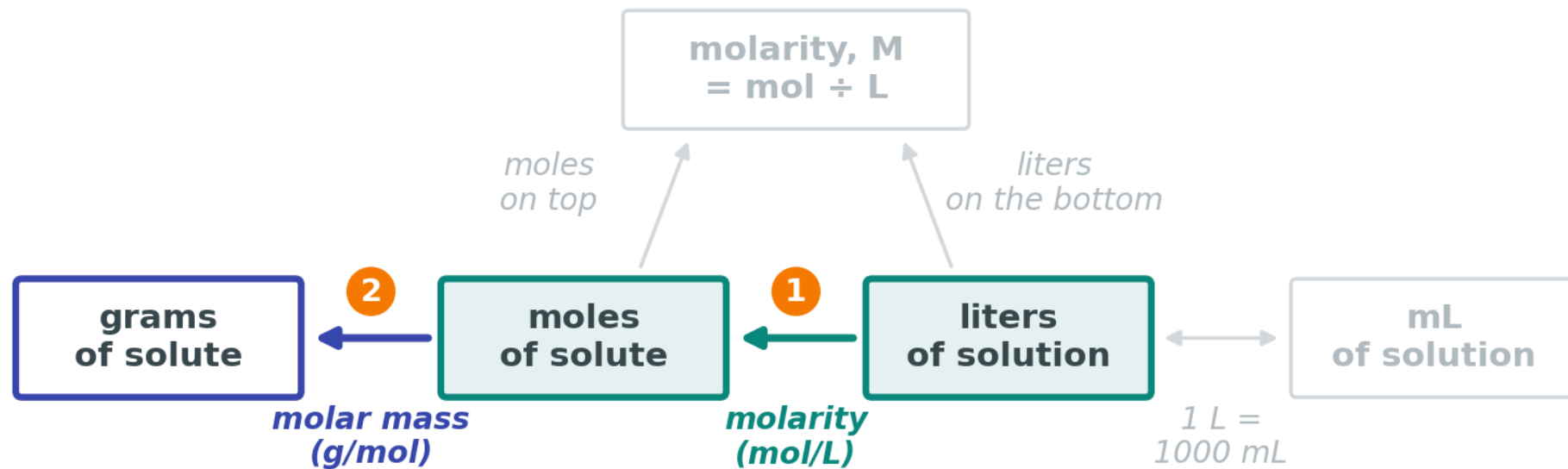
$$1.0 \text{ L soln} \times \frac{0.10 \text{ mol NaOH}}{1 \text{ L soln}} \times \frac{40.00 \text{ g NaOH}}{1 \text{ mol NaOH}} = 4.0 \text{ g NaOH}$$

 **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



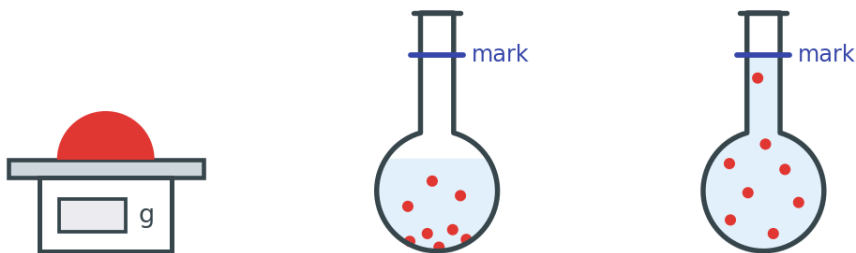
🧠 **Sense check:** Liters in, grams out: the molarity reaches moles, then the molar mass reaches grams. ✓

✓ Worked example 5: making the solution

0.10 mol NaOH = 1 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

$$500.0 \text{ mL soln} \times \frac{1 \text{ L soln}}{1000 \text{ mL soln}} \times \frac{0.150 \text{ mol Na}_2\text{SO}_4}{1 \text{ L soln}} \times \frac{142.04 \text{ g Na}_2\text{SO}_4}{1 \text{ mol Na}_2\text{SO}_4} = 10.7 \text{ g (answer B)}$$

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.

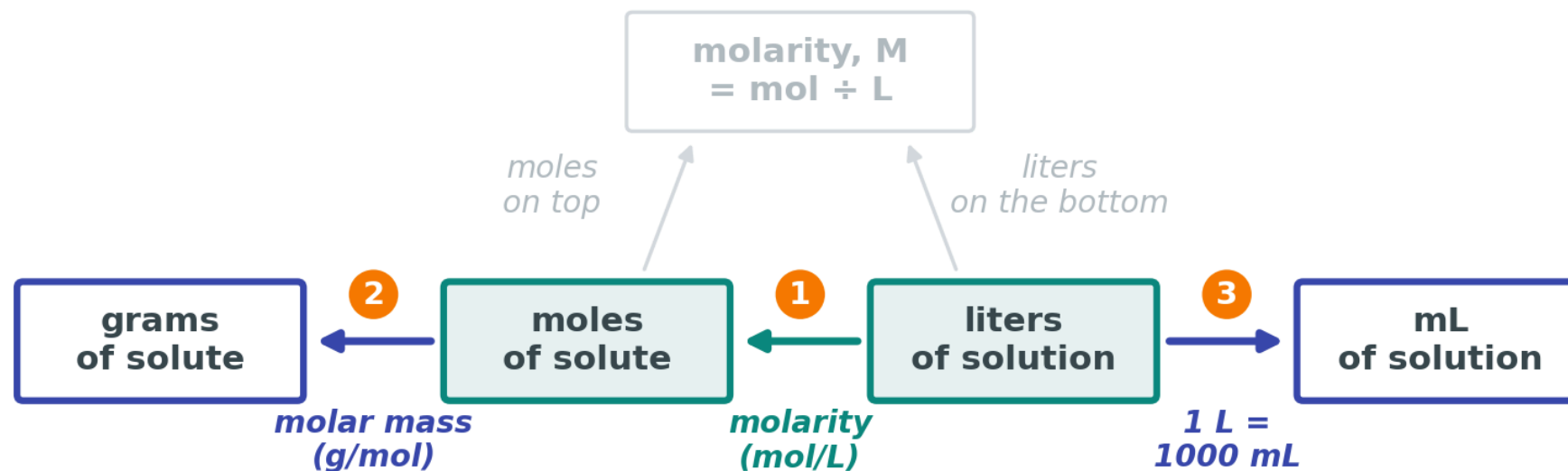
$$\frac{20.6 \text{ g } \text{C}_2\text{H}_6\text{O}}{5.6 \times 10^3 \text{ mL soln}} \times 100 = \mathbf{0.37\% \text{ (m/v)}}$$

 **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

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

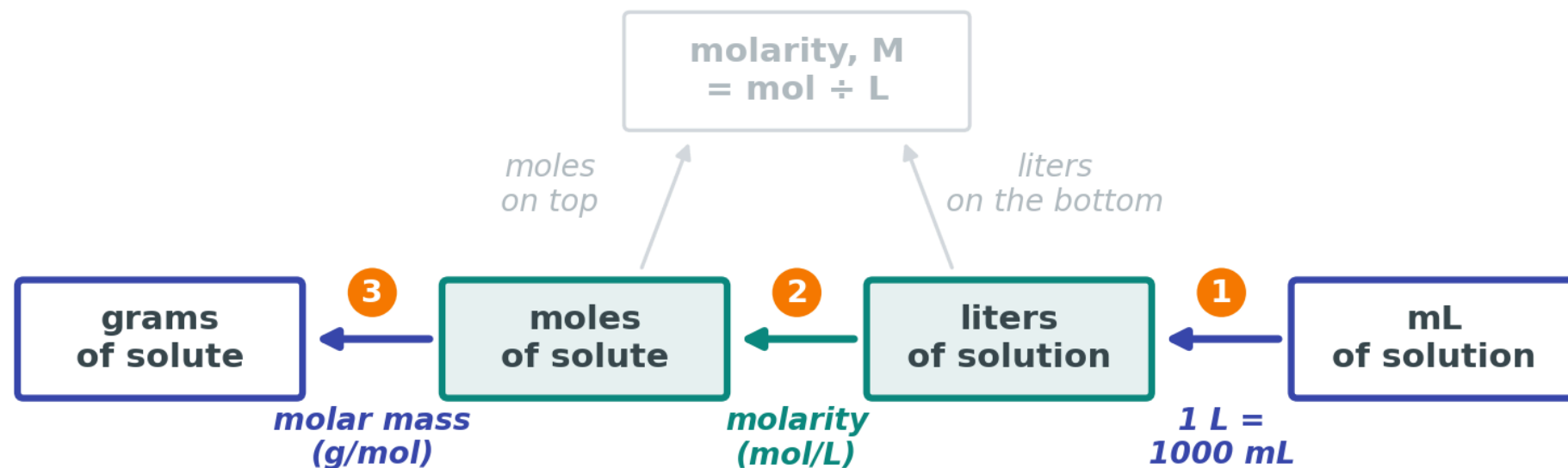
A filled one flask: $0.2500 \times 0.200 \times 159.61 = 7.98$. B flipped the molarity: $1.500 \div 0.200 \times 159.61 = 1.20 \times 10^3$. D stopped at moles: $1.500 \times 0.200 = 0.300 \text{ mol}$. E skipped mL → L: $1500. \times 0.200 \times 159.61 = 4.79 \times 10^4$.

🧠 **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?

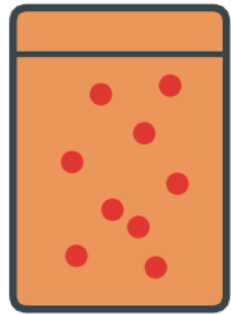
A stock solution is often more concentrated than a procedure needs. Adding water lowers the molarity while the moles of solute stay the same, so $M_1V_1 = M_2V_2$.

2 • 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.

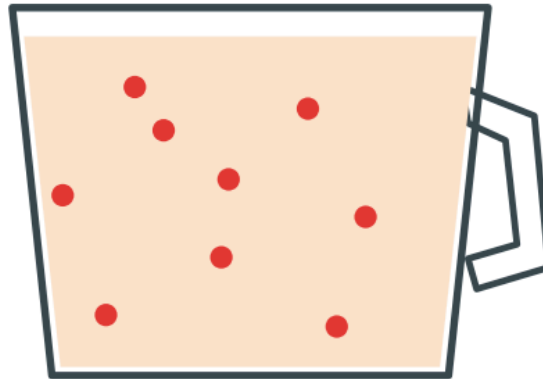
One can makes a whole pitcher

the same flavor bits — spread through much more liquid



one small can of
thick, strong concentrate

+ a pitcher of water

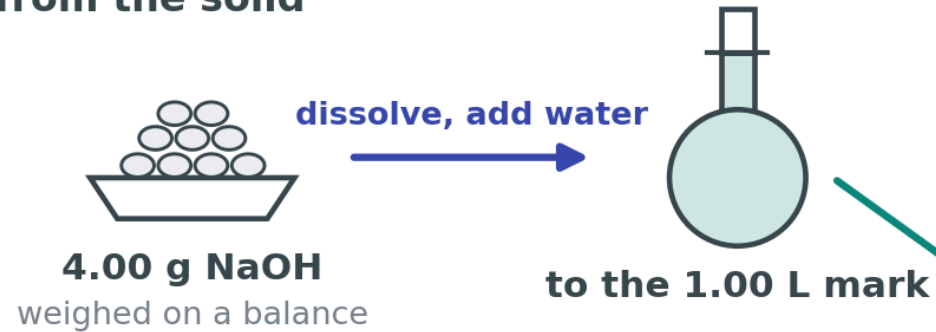


a full pitcher of
pale, mild juice

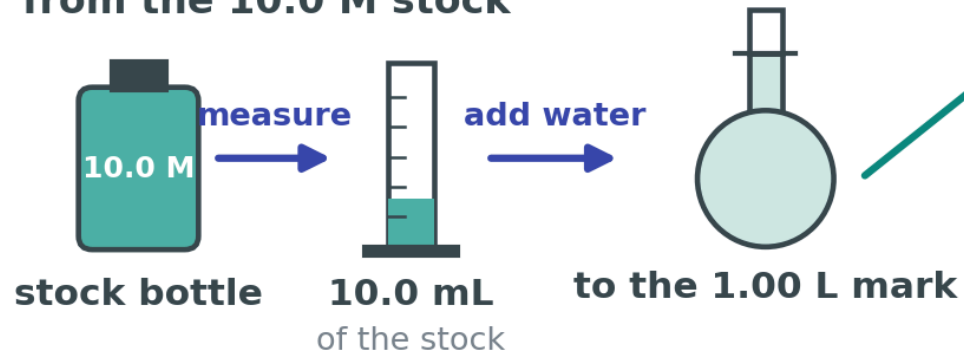
Frozen juice concentrate is thick and strong. Stir one small can into a pitcher of water and it becomes a full, mild drink. Same juice, just more liquid.

A stock bottle skips the weighing

from the solid



from the 10.0 M stock



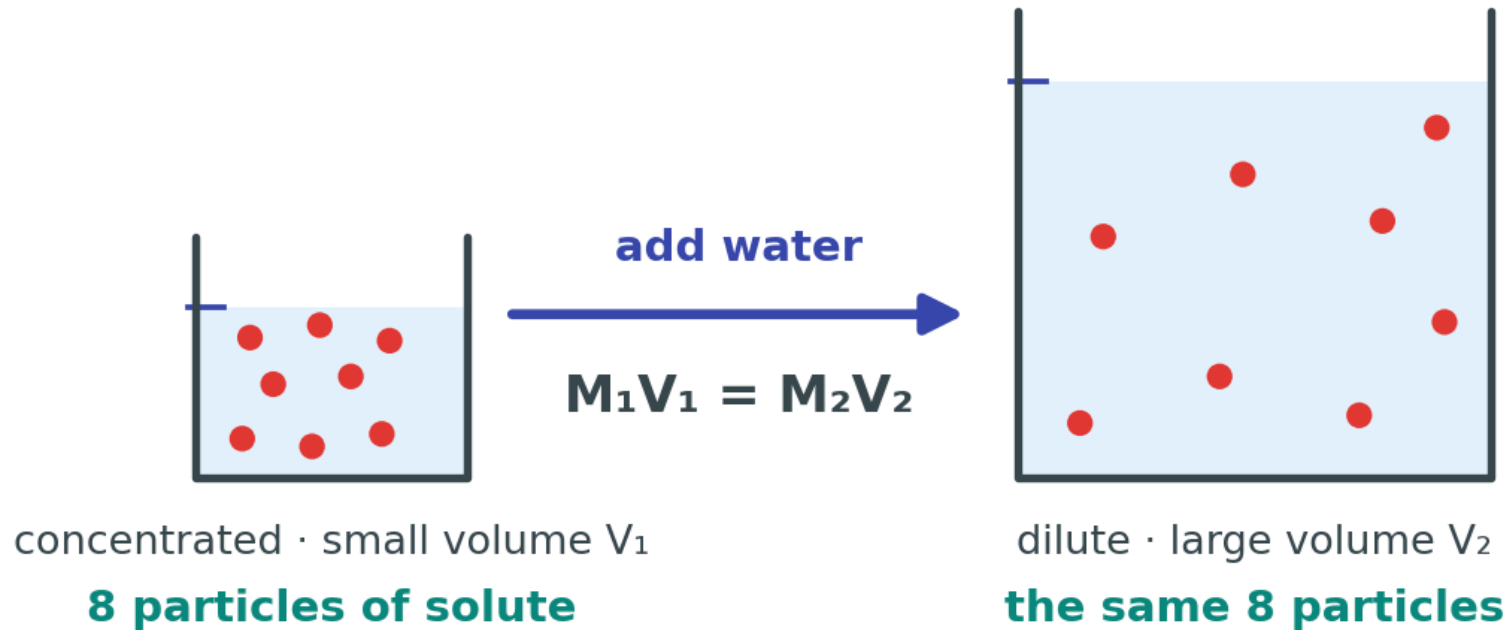
the same flask

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

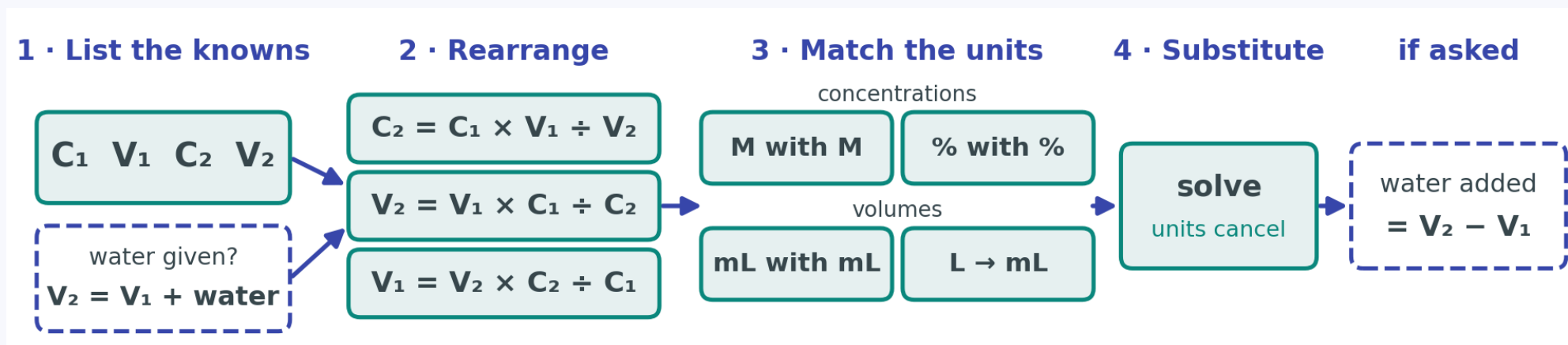
$$C_1V_1 = C_2V_2$$

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

C_1 V_1 C_2 V_2

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

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

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

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

concentrations

M with M

% with %

volumes

mL with mL

L → mL

solve
units cancel

water added
 $= V_2 - V_1$

🧠 **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

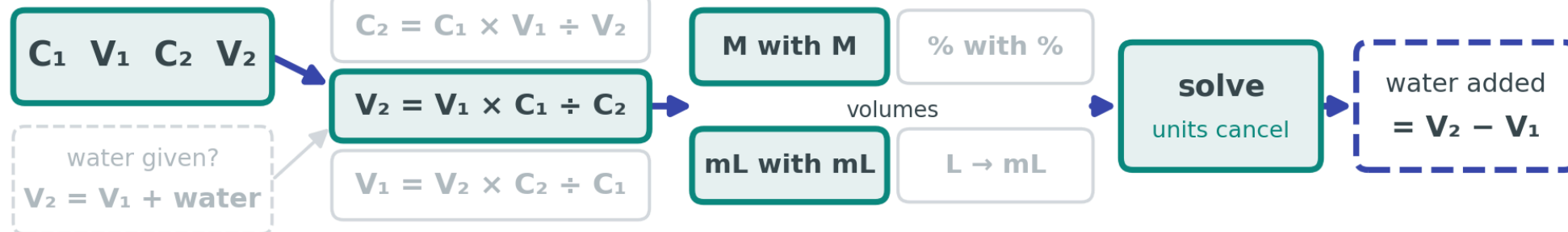
1 · List the knowns

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

if asked

solve
units cancel

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: $28.0 \times 4.00 = 112 \text{ mmol}$ 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: $0.300 \div 6.67 = 0.0450 \text{ M} \checkmark$



Check yourself

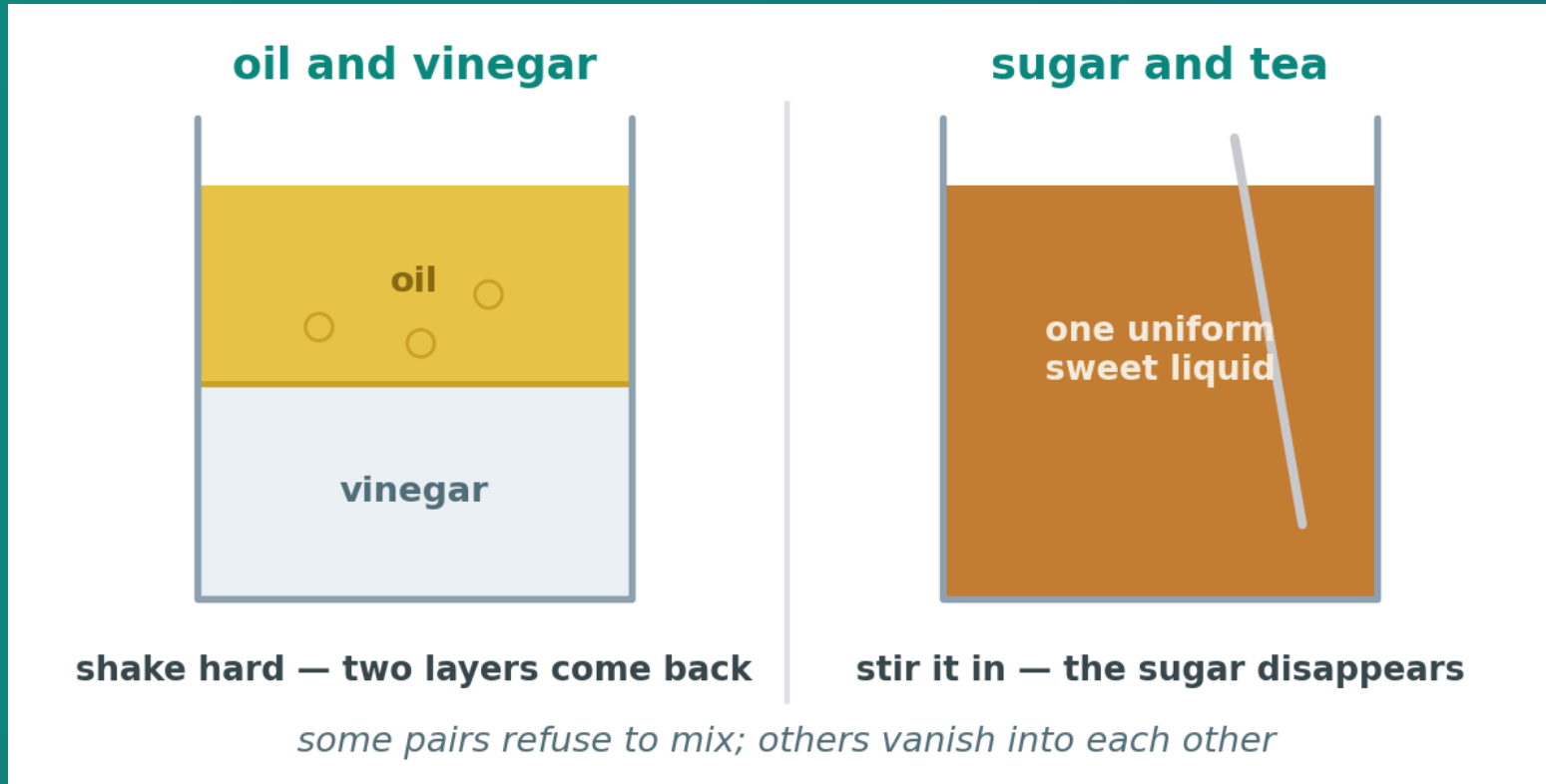
1. A bottle reads **6.00 M NaOH**. Write **$M_1V_1 = M_2V_2$** for making **250. mL of 0.300 M NaOH**, with the three known values filled in. Which volume is larger: the stock you measure, or 250. mL?
2. You dilute a stock and your result for the final volume comes out smaller than the volume you started 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.

3 • Like Dissolves Like

Predict whether a solute dissolves in a given solvent by matching their attractions: polar and ionic solutes dissolve in polar solvents, nonpolar solutes in nonpolar solvents; then explain, from the forces broken and formed, why dissolving can cool or warm the solution.

Why oil won't mix but sugar vanishes



Shake oil and vinegar and they split back into two layers every time. Stir sugar into hot tea and it disappears completely.

Structure decides what dissolves

like dissolves like

a solute dissolves in a solvent whose bonding and attractions are similar to its own

polar family: ionic compounds · molecules with O-H, N-H or a lopsided shape
table salt · water · ammonia

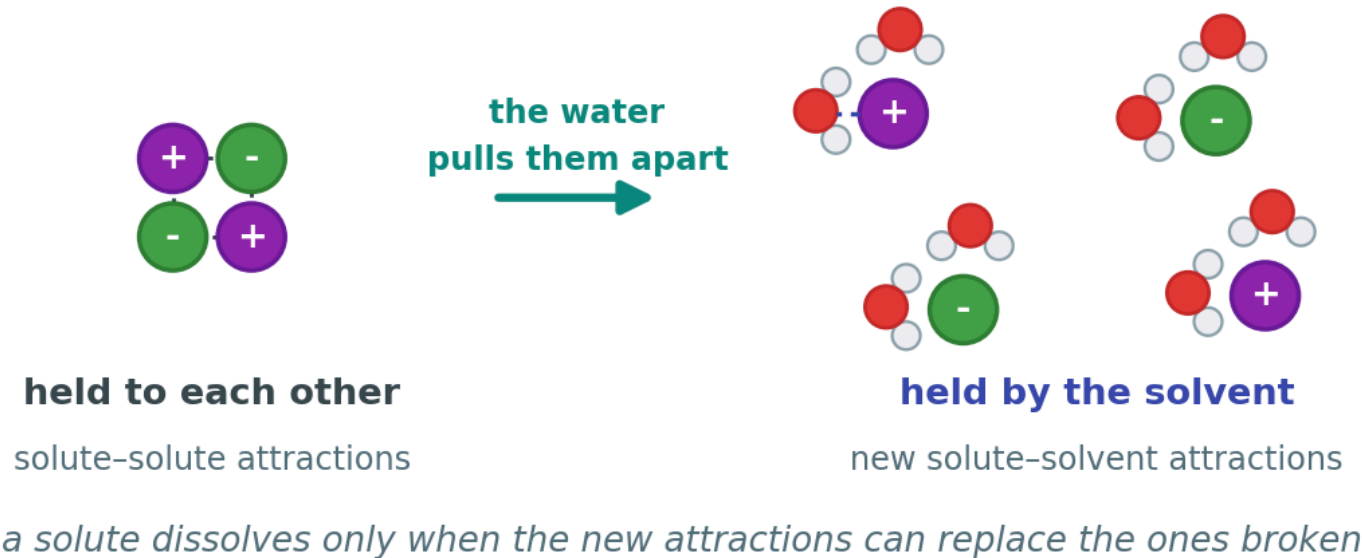
nonpolar family: molecules of only C and H · symmetric molecules
gasoline · vegetable oil · carbon dioxide

Sort baking soda (NaHCO_3) and propane (C_3H_8) into their families.

baking soda: Na^+ and HCO_3^- ions → polar family
propane: only C and H → nonpolar family

A solute dissolves by replacing attractions

dissolving trades one set of attractions for another



The **solute** is what dissolves; the **solvent** is what it dissolves into. A solute dissolves only when new solute-solvent attractions can replace the solute-solute attractions being broken.

Dissolving a salt: forces broken, forces formed

the cost: ionic bonds in the solid break · hydrogen bonds between water molecules break

pulling the solute apart and the solvent apart both take energy

the payback: ion-dipole forces form between each ion and water

ion-dipole force = the attraction between an ion and a polar molecule · forming it releases energy

The **heat of solution** is the difference: energy to separate solute and solvent, less energy released as new attractions form. A molecular solute breaks its intermolecular forces instead of ionic bonds.

Cold packs and hot packs

cost > payback → endothermic · the solution cools

the extra energy is drawn from the solution's own heat · KNO_3 or NH_4NO_3 in water: the cold pack

payback > cost → exothermic · the solution warms

the surplus energy is released as heat · CaCl_2 in water: the hot pack

Dissolving is favored by a drop in energy and by a rise in disorder as the solute spreads out. The disorder is why KNO_3 still dissolves despite the energy cost.

Like dissolves like

like dissolves like		
	polar solvent water	nonpolar solvent oil, hexane
polar / ionic solute salt, sugar	dissolves ✓	stays separate ✗
nonpolar solute grease, wax	stays separate ✗	dissolves ✓

matching families dissolve; a mismatch stays separate

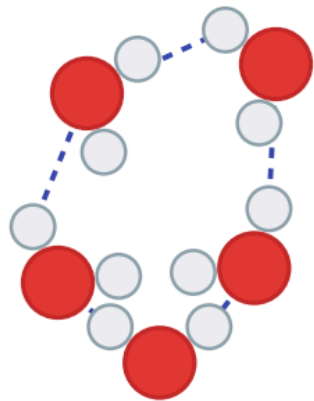
Attractions match only when polarities match. A **polar** or ionic solute dissolves in a polar solvent like water. A **nonpolar** solute dissolves in a nonpolar solvent like oil or hexane.

memory hook: water loves charge, oil loves oil

polar and ionic go with polar · nonpolar goes with nonpolar · a mismatch stays in two layers

Why oil and water separate

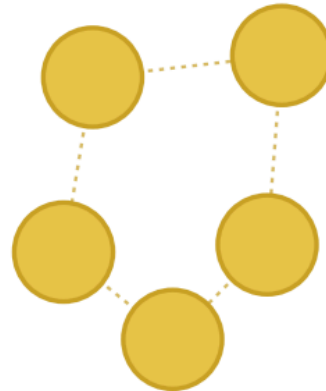
water holds itself; the oil is squeezed out



water clings to water

strong hydrogen bonds

no strong attraction bridges the two — so they split into layers



oil clings to oil

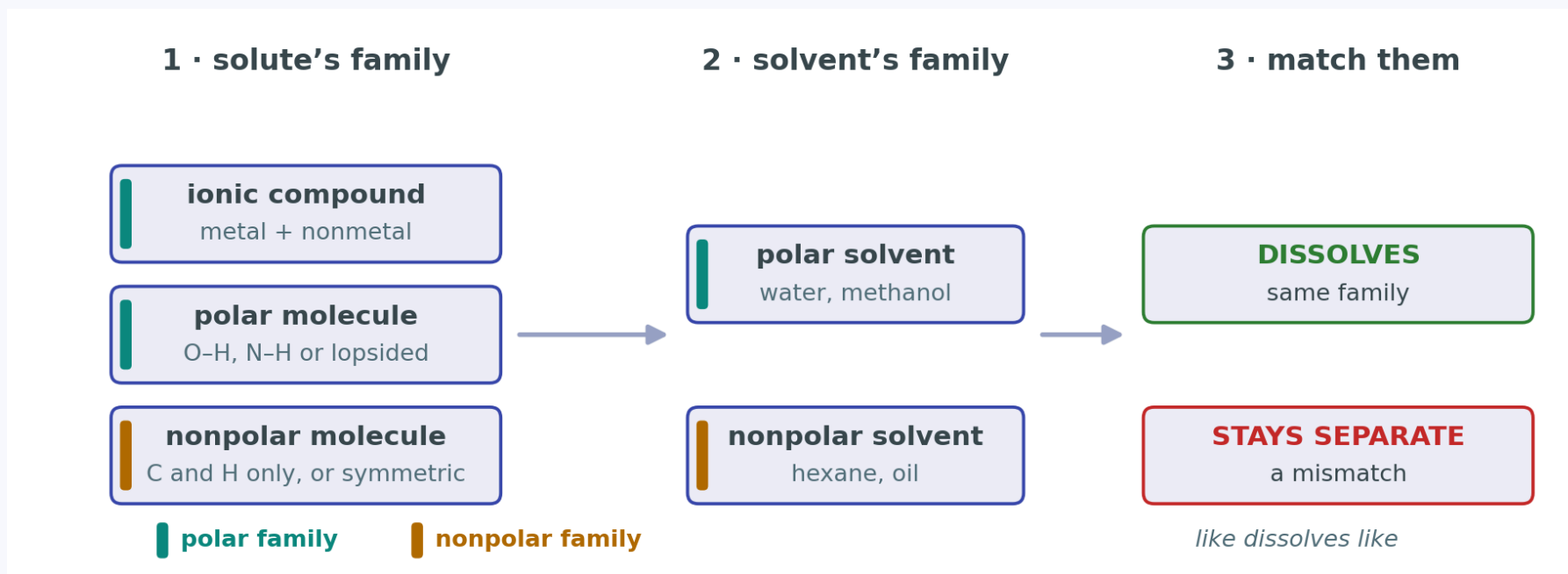
only weak attractions

Water molecules hold one another with strong hydrogen bonds. A nonpolar oil molecule attracts only weakly, too weak to enter that network. So the oil is pushed out into its own layer.



The method

1. **Find the solute's polarity.** Ionic or polar, or nonpolar?
2. **Find the solvent's polarity.** Polar like water, or nonpolar like oil or hexane?
3. **Match them.** Same family dissolves. A mismatch stays separate.





Guided example: carbon tetrachloride in hexane

a few drops of carbon tetrachloride (CCl_4) added to hexane (C_6H_{14})

solute: CCl_4 , four polar C-Cl bonds · solvent: hexane · wanted: dissolves or stays separate?

Carbon tetrachloride and hexane are both clear liquids. Do they dissolve in one another?

Sort each liquid into its family first, then match the families.

✓ Guided example: solution

solute: carbon tetrachloride (CCl_4) · solvent: hexane (C_6H_{14})

Step 1 · Find the solute's polarity

$\Delta\text{EN}(\text{C}-\text{Cl}): 3.16 - 2.55 = \mathbf{0.61}$ · polar bonds · symmetric tetrahedron → **nonpolar solute**

Step 2 · Find the solvent's polarity

C_6H_{14} : only C and H → **nonpolar solvent**

Step 3 · Match them

nonpolar solute · nonpolar solvent → **same family, they dissolve**

🧠 **Sense check:** CCl_4 's polar bonds cancel by shape. Old and new dispersion forces match, so they mix freely.

✓ Guided example: the route on the map

carbon tetrachloride (CCl_4) in hexane (C_6H_{14})

nonpolar solute · nonpolar solvent · found: they dissolve in one another

1 · solute's family

CCl_4 : symmetric

ionic compound
metal + nonmetal

polar molecule
O-H, N-H or lopsided

nonpolar molecule
C and H only, or symmetric

polar family

nonpolar family

2 · solvent's family

hexane

polar solvent
water, methanol

nonpolar solvent
hexane, oil

3 · match them

they mix

DISSOLVES
same family

STAYS SEPARATE
a mismatch

like dissolves like

 **Sense check:** Both liquids sit in the nonpolar family. Same family: they mix. ✓



Practice 1: sodium iodide

sodium iodide, NaI

liquids on the shelf: hexane (C_6H_{14}) · carbon tetrachloride (CCl_4) · methanol (CH_3OH)

Which liquid dissolves sodium iodide?

- A. Hexane
- B. Methanol
- C. Carbon tetrachloride
- D. None of them: an ionic solid dissolves only in water

✓ Practice 1: answer B

sodium iodide, NaI: ionic · methanol, CH₃OH: polar, an O-H group
ionic solute · polar solvent · same family → it dissolves

NaI: Na⁺ and I⁻ ions → polar family · CH₃OH: an O-H group → polar solvent
→ it dissolves (answer B)

A paired an ionic solid with a nonpolar liquid, as if opposite polarities attract. C read CCl₄'s polar C-Cl bonds as a polar molecule; its symmetric shape cancels them. D treated water as the only polar solvent; methanol's O-H group makes it polar too.

🧠 **Sense check:** Methanol's O-H end turns toward the ions, as water's does. Water is not the only polar solvent. ✓

✓ Practice 1: the route on the map

sodium iodide (NaI) in methanol (CH₃OH)
ionic solute · polar solvent · found: it dissolves

1 · solute's family

NaI: Na⁺ and I⁻

ionic compound
metal + nonmetal

polar molecule
O-H, N-H or lopsided

nonpolar molecule
C and H only, or symmetric

polar family

nonpolar family

2 · solvent's family

methanol: O-H

polar solvent
water, methanol

nonpolar solvent
hexane, oil

3 · match them

it dissolves

DISSOLVES
same family

STAYS SEPARATE
a mismatch

like dissolves like

🧠 **Sense check:** Ionic solute, polar solvent: one family. Any polar solvent, not only water, takes an ionic solute. ✓



Worked example 1: table salt in water

table salt (NaCl) stirred into water

ionic solid · 58.44 g/mol · water is a polar solvent

Table salt is an ionic solid, and water is a polar solvent. Predict whether the salt dissolves, and why.

✓ Worked example 1: solution

table salt (NaCl) stirred into water
ionic solid · 58.44 g/mol · water is a polar solvent

Step 1 · Find the solute's polarity

Step 2 · Find the solvent's polarity

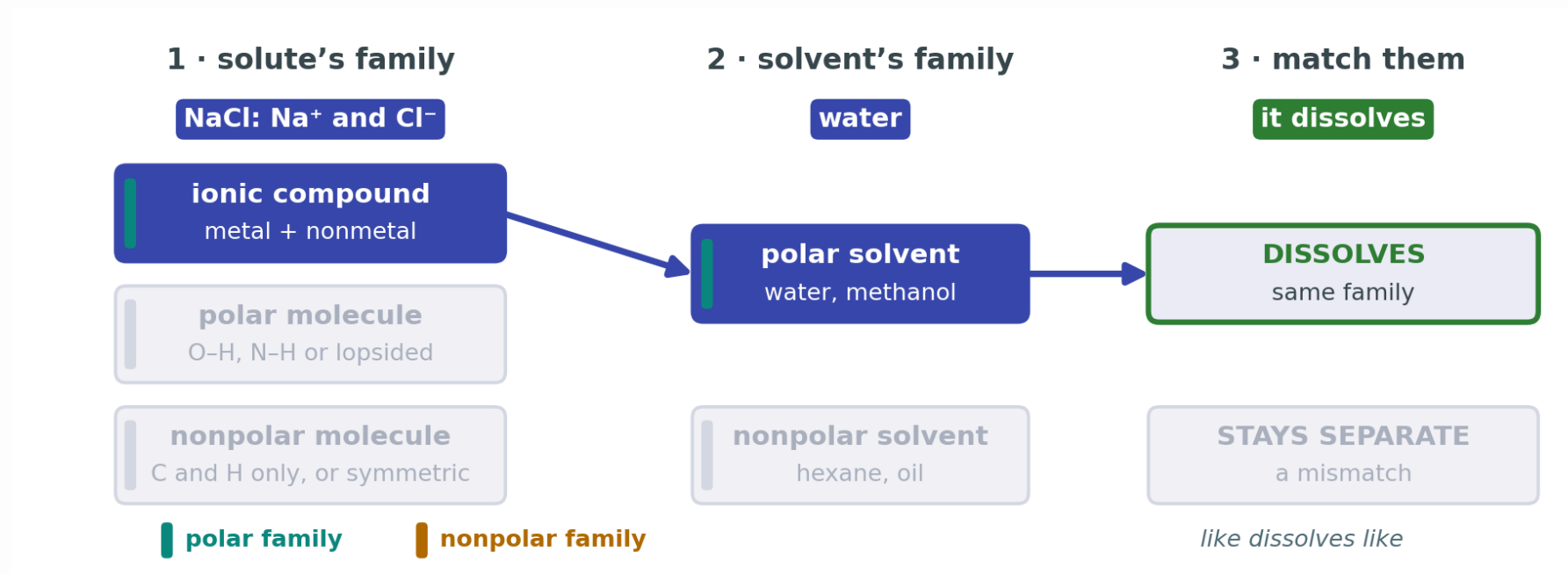
Step 3 · Match them

ionic solute · polar solvent → **same family, it dissolves**

🧠 **Sense check:** The new ion-water attractions are as strong as the ones inside the salt, so the crystal comes apart and dissolves.

✓ Worked example 1: the route on the map

table salt (NaCl) stirred into water
ionic solute · polar solvent · found: it dissolves



🧠 **Sense check:** An ionic solute sits in the polar family, the same family as water. Same family: it dissolves. ✓



Worked example 2: cooking grease in water

cooking grease rinsed with cold water

nonpolar solute · solvent: water, polar

Grease is nonpolar, and you try to rinse a greasy pan with plain cold water.

A common first answer: water dissolves so much that enough of it must wash the grease away. Test it against the method.

✓ Worked example 2: solution

cooking grease rinsed with cold water
nonpolar solute · solvent: water, polar

A common first answer

enough water washes the grease away?
water dissolves polar and ionic things, not nonpolar grease ✕

Step 1 · Find the solute's polarity

Step 2 · Find the solvent's polarity

Step 3 · Match them

nonpolar solute · polar solvent → **a mismatch, it stays separate**

 **Sense check:** The attraction between grease and water is far weaker than water's attraction to itself. The grease beads up and stays put, however much water runs over it.

✓ Worked example 2: the route on the map

cooking grease rinsed with cold water
nonpolar solute · polar solvent · found: it stays separate

1 · solute's family

grease: mostly C-H

ionic compound
metal + nonmetal

polar molecule
O-H, N-H or lopsided

nonpolar molecule
C and H only, or symmetric

polar family

nonpolar family

2 · solvent's family

water

polar solvent
water, methanol

nonpolar solvent
hexane, oil

3 · match them

two layers

DISSOLVES
same family

STAYS SEPARATE
a mismatch

like dissolves like

🧠 **Sense check:** Grease sits in the nonpolar family, water in the polar family. A mismatch leaves two layers. ✓

Take-home: water is not a universal solvent

salt · sugar · vinegar: dissolve in water

polar and ionic solutes: water's charged ends attract them

oil · grease · wax: do not dissolve in water

nonpolar solutes: water bonds to itself and leaves them out

Water dissolves polar and ionic solutes, not everything. A nonpolar solute cannot replace the strong hydrogen bonds between water molecules, so it stays separate however much water is present.

Your turn: candle wax in water

a chip of candle wax dropped into water
wax is nonpolar · water is a polar solvent

step	question	answer
1 · solute polarity	polar/ionic, or nonpolar?	wax is _____
2 · solvent polarity	polar or nonpolar?	water is _____
3 · match them	same family, or a mismatch?	_____

Fill the three cells, then read off the result.

candle wax + water → stays separate
nonpolar solute · polar solvent · a mismatch: the wax needs a nonpolar solvent

⚠ Where this goes wrong

Treating water as a universal solvent. Water dissolves many things, but only polar and ionic ones. A nonpolar solute like oil or grease is attracted to water far more weakly than water is to itself, so no amount of water dissolves it.

Arguing that opposite polarities attract. Opposite charges attract, plus to minus. Solubility does not work that way. It follows like dissolves like: a nonpolar solute needs a nonpolar solvent, never a polar one.

Expecting stirring to overcome a mismatch. Stirring and waiting change how fast a solute dissolves, never whether it dissolves. A polarity mismatch stays undissolved however long you stir.

Reading a big carbon molecule as automatically nonpolar. Sugar and glycerol are built on carbon, yet their many -OH groups are polar and hydrogen-bond with water. Count the polar groups before calling a molecule nonpolar.



Practice 2

hexane, C_6H_{14} : a nonpolar solvent

candidates: table salt · glycerol · naphthalene (mothballs)

A chemist needs a solute that dissolves readily in hexane. Which candidate works best, and why?

- A. Table salt: an ionic solid, and ionic solids dissolve in almost any liquid
- B. Glycerol, $\text{C}_3\text{H}_8\text{O}_3$: built on carbon, so it counts as nonpolar
- C. Naphthalene, C_{10}H_8 : a nonpolar solute, the same family as hexane
- D. Any of the three, as long as the mixture is stirred long enough

✓ Practice 2 answer: C

hexane (nonpolar) + naphthalene (nonpolar) → dissolves · answer C
same family · the other candidates are ionic or polar

A treats ionic solids as dissolving anywhere; ions need a polar solvent whose charged ends can surround them, and hexane has none. B judged glycerol by its carbon atoms; its three -OH groups make it polar, so it mixes with water, not hexane. D counts on stirring, which changes only the rate, never whether a mismatch dissolves.

 **Sense check:** Match the families first: nonpolar hexane takes the nonpolar naphthalene and leaves the salt and glycerol behind.

✓ Practice 2: the route on the map

naphthalene ($C_{10}H_8$) in hexane (C_6H_{14})
nonpolar solute · nonpolar solvent · found: it dissolves

1 · solute's family

naphthalene: C and H

ionic compound
metal + nonmetal

polar molecule
O-H, N-H or lopsided

nonpolar molecule
C and H only, or symmetric

polar family

nonpolar family

2 · solvent's family

hexane

polar solvent
water, methanol

nonpolar solvent
hexane, oil

3 · match them

it dissolves

DISSOLVES
same family

STAYS SEPARATE
a mismatch

like dissolves like

🧠 **Sense check:** Only C and H in both: the nonpolar family on both sides, so naphthalene dissolves. ✓



Worked example 3: sugar in water

table sugar ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) stirred into water

carbon-heavy molecule · 342.30 g/mol · solvent: water, polar

Sugar is a large molecule built on a carbon framework, which makes it look nonpolar. Predict whether it dissolves in water, and why.

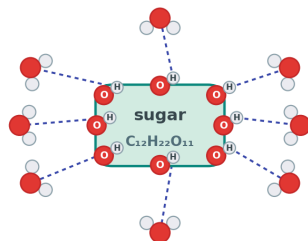
✓ Worked example 3: the polar groups

table sugar ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) stirred into water
carbon-heavy molecule · 342.30 g/mol · solvent: water, polar

Step 1 · Find the solute's polarity

The carbon framework is nonpolar, but the molecule carries eight polar -OH groups.

every -OH hydrogen-bonds to a water molecule



big carbon framework, but covered in polar -OH groups → dissolves in water

 **Sense check:** The carbon count is a distraction. The eight polar -OH groups decide it: sugar is polar.

Step 2 · Find the solvent's polarity

Water is polar, and every one of those -OH groups can hydrogen-bond to it.

✓ Worked example 3: the match

table sugar ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$): polar, covered in -OH groups
solvent: water, polar

Step 3 · Match them

polar solute · polar solvent → **same family, it dissolves**

🧠 **Sense check:** Do not judge a molecule by its carbon count. Sugar's eight polar -OH groups make it polar, so it dissolves in water. The tea turns sweet.

✓ Worked example 3: the route on the map

table sugar ($C_{12}H_{22}O_{11}$) stirred into water
eight -OH groups: polar solute · polar solvent · found: it dissolves

1 · solute's family

sugar: 8 -OH

ionic compound
metal + nonmetal

polar molecule
O-H, N-H or lopsided

nonpolar molecule
C and H only, or symmetric

polar family

nonpolar family

2 · solvent's family

water

polar solvent
water, methanol

nonpolar solvent
hexane, oil

3 · match them

it dissolves

DISSOLVES
same family

STAYS SEPARATE
a mismatch

like dissolves like

 **Sense check:** The -OH groups put sugar in the polar family with water, carbon framework and all. ✓

Practice 3: two vitamins

vitamin A, $C_{20}H_{30}O$: one $-OH$ group · vitamin C, $C_6H_8O_6$: four $-OH$ groups
solvents: water · body fat

Which vitamin dissolves in water, and which in body fat?

- A. Vitamin C in water; vitamin A in body fat
- B. Both in water: each has at least one polar $-OH$ group
- C. Vitamin A in water; vitamin C in body fat
- D. Both in body fat: both are built on carbon

✓ Practice 3: answer A

vitamin C, $C_6H_8O_6$: four -OH on six carbons → polar → water

vitamin A, $C_{20}H_{30}O$: one -OH on twenty carbons → nonpolar → body fat

vitamin C: $6 \text{ C} \div 4 \text{ -OH} = 1.5 \text{ C per -OH}$ · vitamin A: $20 \text{ C} \div 1 \text{ -OH} = 20 \text{ C per -OH}$

→ **answer A**

B counted any -OH as enough; one -OH cannot carry a 20-carbon chain into water. C swapped the two families. D judged both by the carbon framework; vitamin C's four -OH groups make it polar.

 **Sense check:** Vitamin C has one -OH per 1.5 carbons, the same share as sugar, so it dissolves in water. Vitamin A's lone -OH is swamped by its carbon chain, so the body stores it in fat. ✓

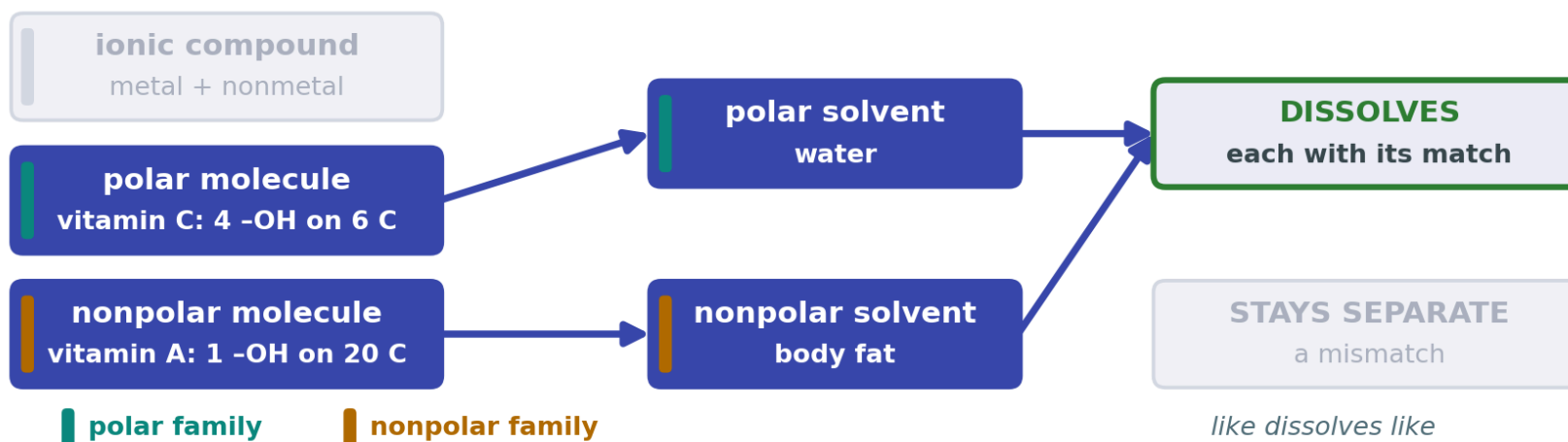
✓ Practice 3: the route on the map

vitamin C ($C_6H_8O_6$) and vitamin A ($C_{20}H_{30}O$)
 vitamin C: polar → water · vitamin A: nonpolar → body fat

1 · solute's family

2 · solvent's family

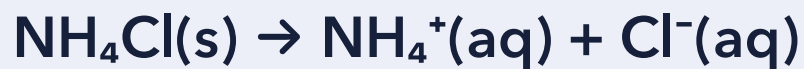
3 · match them



🧠 **Sense check:** Two routes, one rule: the share of -OH groups sets each vitamin's family. ✓



Practice 4: a salt that cools the water



ammonium chloride dissolving in water · the solution turns cold

Which statement explains the cooling?

- A. Exothermic: breaking the ionic bonds in the solid releases energy
- B. Endothermic: forming ion-dipole forces with water absorbs energy
- C. Exothermic: the solution cools because dissolving releases heat
- D. Endothermic: forming ion-dipole forces releases more than separating costs
- E. Endothermic: separating ions and water costs more than ion-dipole forces release

✓ Practice 4: answer E



cost: ionic bonds and water's hydrogen bonds break · payback: ion-dipole forces form

cost > payback → **endothermic: the solution cools (answer E)**

A reversed bond breaking: pulling ions apart always costs energy. B reversed bond forming: forming ion-dipole forces releases energy. C read the cooling backwards: dissolving draws heat from the water, so the process absorbs it. D kept the label but flipped the comparison; a larger payback would warm the solution.

 **Sense check:** A solution that turns cold has paid more to separate than it got back. The heat came from the water itself. ✓



Practice 5

copper(II) sulfate (CuSO_4) · erythritol ($\text{C}_4\text{H}_{10}\text{O}_4$, an -OH on every carbon) · β -carotene ($\text{C}_{40}\text{H}_{56}$)

the mixture is shaken with water and hexane, which settle into two layers

Where does each solute end up?

- A. All three in the water layer
- B. CuSO_4 in the water layer; erythritol and β -carotene in the hexane layer
- C. CuSO_4 and erythritol in the hexane layer; β -carotene in the water layer
- D. CuSO_4 and erythritol in the water layer; β -carotene in the hexane layer

✓ Practice 5 answer: D

water layer: CuSO_4 (ionic) and erythritol (polar) · hexane layer: β -carotene (nonpolar)
· answer D

ionic and polar solutes with the polar solvent · nonpolar solute with the nonpolar solvent

CuSO_4 is ionic, so it belongs with polar water. Erythritol is built on carbon, but an $-\text{OH}$ on every carbon makes it polar: water again. β -Carotene holds only C and H, so it is nonpolar: hexane.

A treated water as a universal solvent; it leaves nonpolar β -carotene out. B judged erythritol by its carbon skeleton instead of its $-\text{OH}$ groups. C swapped the families, as if opposite polarities attract.

 **Sense check:** Sort each solute on its own: ionic or polar goes with water, nonpolar goes with hexane. Shaking changes only how fast they sort.

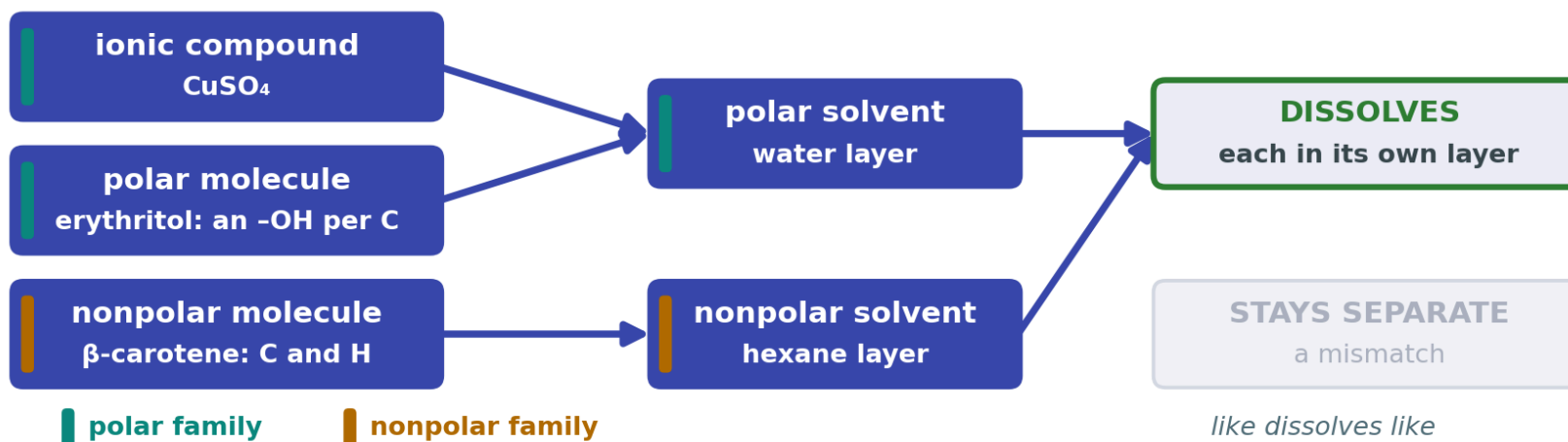
✓ Practice 5: the route on the map

$\text{CuSO}_4 \cdot \text{erythritol (C}_4\text{H}_{10}\text{O}_4) \cdot \beta\text{-carotene (C}_{40}\text{H}_{56})$ in water and hexane
 ionic and polar solutes $\rightarrow$ water layer · nonpolar solute $\rightarrow$ hexane layer

1 · solute's family

2 · solvent's family

3 · match them



🧠 **Sense check:** Three solutes, three routes: each one follows its own family into the layer that matches. ✓

Extra practice 1

limonene, the oil pressed from orange peel

shaken with water: two separate layers · shaken with hexane (C_6H_{14}): one clear layer, in any amount

What do these two results show about limonene?

- A. It is nonpolar, the same family as hexane
- B. It is polar: opposite polarities attract, so it mixes with hexane
- C. Nothing about polarity: the separate layer shows only that it is less dense than water
- D. Nothing about polarity: shaken longer, it would mix into water too

✓ Extra practice 1 answer: A

limonene: one layer with hexane · two layers with water → nonpolar · answer A
it mixes with the nonpolar solvent and stays out of the polar one

mixes with hexane → like hexane · separate from water → unlike water
→ **nonpolar (answer A)**

B ran the rule backward, as if opposite polarities attract; a polar liquid would mix with water, not hexane. C read the layers as density alone; ethanol is also less dense than water, yet it mixes with water in any amount. D counted on shaking, which changes only how fast liquids mix, never whether a mismatch mixes.

 **Sense check:** Read the rule in reverse: a liquid shares the polarity of the solvent it mixes with. Limonene, $C_{10}H_{16}$, holds only C and H, which fits. ✓

Extra practice 2

potassium bromide (KBr) · urea, $(\text{NH}_2)_2\text{CO}$ · cholesterol, $\text{C}_{27}\text{H}_{45}\text{OH}$
the three solids are shaken with water and hexane, which settle into two layers

Which solute or solutes end up in the hexane layer?

- A. KBr and urea
- B. Urea and cholesterol
- C. None: cholesterol's -OH group carries it into the water
- D. Cholesterol only
- E. All three, once the mixture is shaken long enough

✓ Extra practice 2 answer: D

hexane: cholesterol (nonpolar) · water: KBr (ionic), urea (polar) · answer D
ionic and polar solutes with the polar solvent · nonpolar solute with the nonpolar solvent

KBr: K^+ and Br^- ions $\rightarrow$ water · urea: two $-\text{NH}_2$ and a $\text{C}=\text{O}$ $\rightarrow$ polar $\rightarrow$ water ·

cholesterol: one $-\text{OH}$ on 27 C $\rightarrow$ nonpolar $\rightarrow$ **hexane (answer D)**

A swapped the families, as if opposite polarities attract; ions need a polar solvent. B judged urea by its carbon atom; its N-H and C=O groups hydrogen-bond with water. C counted any $-\text{OH}$ as enough; one $-\text{OH}$ cannot carry 27 carbons into water. E counted on shaking, which changes only how fast the solutes sort, not where they end up.

 **Sense check:** Weigh the polar groups against the carbon framework: urea is almost all polar groups, cholesterol almost all C-H. ✓



Extra practice 3

0.0800 mol xylitol = 1 L of solution

xylitol, $\text{C}_5\text{H}_7(\text{OH})_5$: 152.15 g/mol · solvents on hand: water · hexane (C_6H_{14})

A chemist makes **200.0 mL of 0.0800 M xylitol** solution. Which solvent does the job, and what mass of xylitol, in grams, goes in?

- A. Hexane · 2.43
- B. Water · 2.43
- C. Water · 0.0160
- D. Water · 2.43×10^3
- E. Water · 1.05×10^{-4}

✓ Extra practice 3 answer: B

xylitol, $\text{C}_5\text{H}_7(\text{OH})_5$: five -OH groups $\rightarrow$ polar $\rightarrow$ water

given: 200.0 mL of 0.0800 M (152.15 g/mol) · wanted: solvent, g xylitol

Three conversion factors are needed.

$$200.0 \text{ mL soln} \times \frac{1 \text{ L soln}}{1000 \text{ mL soln}} \times \frac{0.0800 \text{ mol}}{1 \text{ L soln}} \times \frac{152.15 \text{ g}}{1 \text{ mol}} = 2.43 \text{ g, in water (answer B)}$$

A judged xylitol by its carbons; five -OH groups make it polar. C stopped at moles: $0.2000 \times 0.0800 = 0.0160$. D skipped mL $\rightarrow$ L: $200.0 \times 0.0800 \times 152.15 = 2.43 \times 10^3$. E divided by the molar mass: $0.0160 \div 152.15 = 1.05 \times 10^{-4}$.

🧠 **Sense check:** 0.0160 mol at about 150 g/mol is about 2.4 g. Its -OH groups pull it into water, as sugar's do. ✓



Check yourself

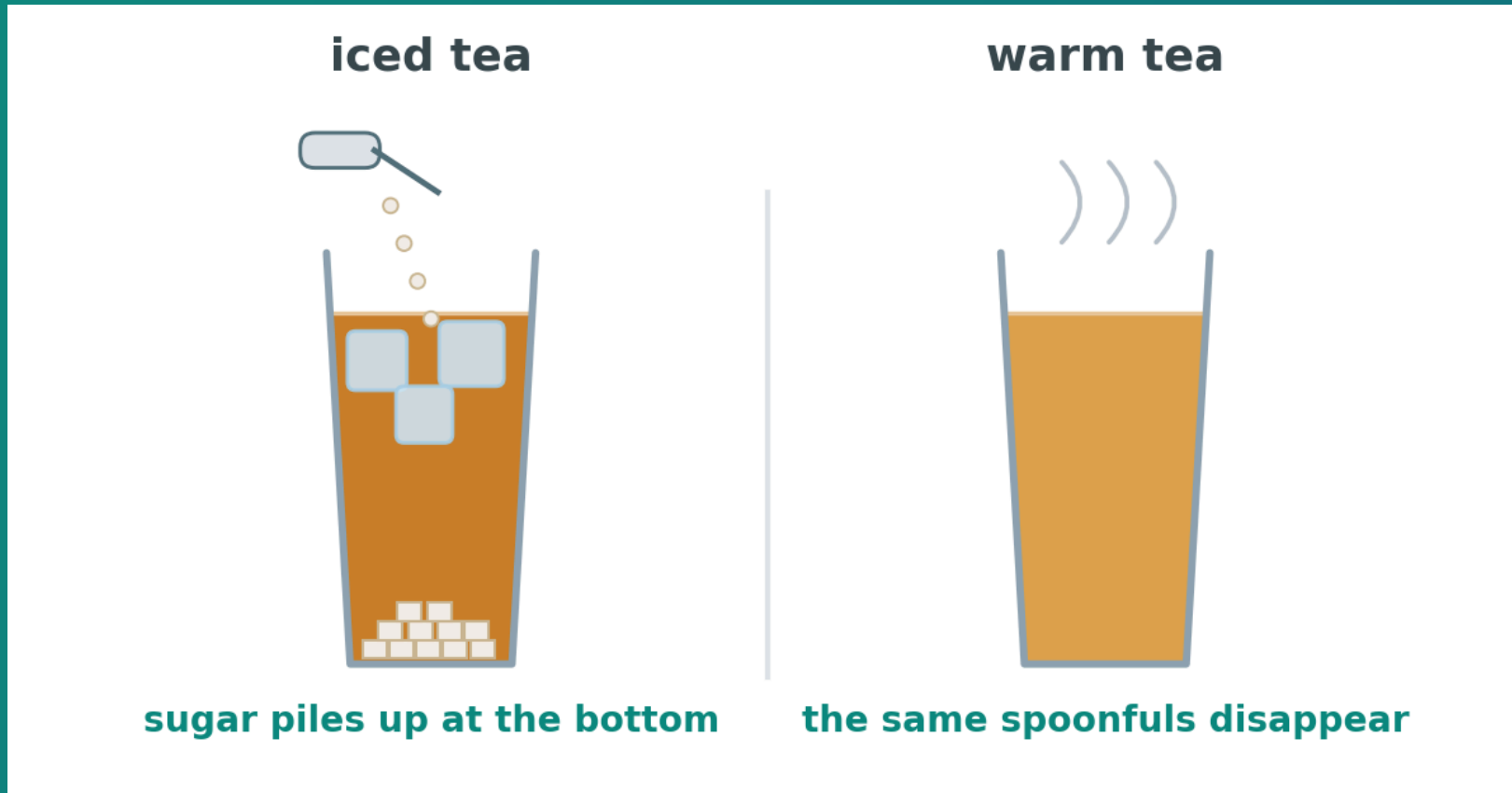
1. Iodine (I_2) is a nonpolar solid. It barely colors water but turns hexane deep purple. Which solvent dissolves it, and why?
2. Rubbing alcohol mixes with water in any amount. What does that tell you about its polarity?
3. Lithium chloride dissolves in water and the solution turns warm. Is dissolving it endothermic or exothermic, and which is larger: the energy to separate the ions and the water molecules, or the energy released as ion-dipole forces form?

Some solutes keep dissolving in a solvent until the solvent can hold no more. That limit, and how temperature shifts it, is the next question about solutions.

4 · Saturation & Solubility Curves

Read a solubility curve to find how much solute dissolves at a given temperature, classify a solution as unsaturated, saturated, or supersaturated, and find how much crystallizes when it cools.

When sugar stops dissolving



Stir spoon after spoon of sugar into iced tea. At first it disappears. Past a point it just piles at the bottom, and warm tea takes far more.

Every solute has a limit

solubility = the maximum solute that dissolves in a fixed amount of water at a given temperature

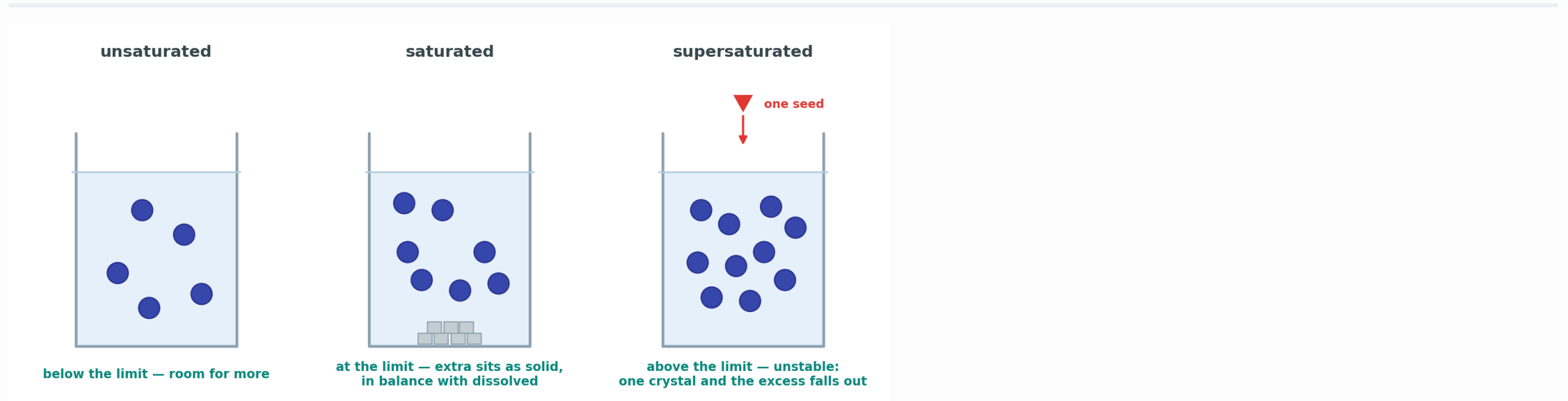
measured in grams of solute per 100 g of water, at a stated temperature

KNO_3 at 20 °C: 32 g per 100 g water

stir in more than 32 g and the extra cannot dissolve: it stays as solid

A fixed amount of water at a fixed temperature dissolves only so much solute. That ceiling, the solute's **solubility**, exists because dissolving and crystallizing reach a balance: particles return as fast as they leave.

Unsaturated, saturated, supersaturated



Below the limit: **unsaturated**, more dissolves. At the limit: **saturated**, solid in balance with dissolved. Past it, after slow cooling with no seed crystal: **supersaturated**. One added crystal drops the excess out.

memory hook: UNDER · AT · OVER the limit

unsaturated · saturated · supersaturated ("super" = over, all of it dissolved)

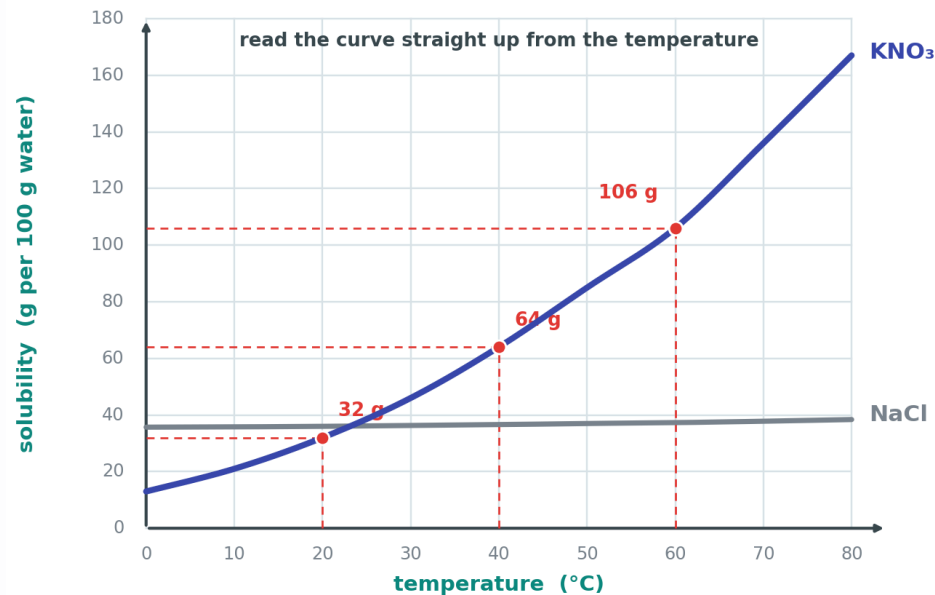
Three factors set the limit

structure · temperature · pressure (gases only)

**like dissolves like · most solids dissolve more when hot, gases less · more gas pressure above the liquid,
more gas dissolves**

The solute and solvent, the temperature, and for a gas the pressure fix the limit. Warm water holds less dissolved oxygen, so fish in a warm pond can run short of it.

Reading the solubility curve



Warm water jostles the crystal apart faster than particles re-stick, so the KNO₃ curve climbs steeply while NaCl stays nearly flat. A warm gas molecule instead escapes the liquid more easily: gases dissolve less.

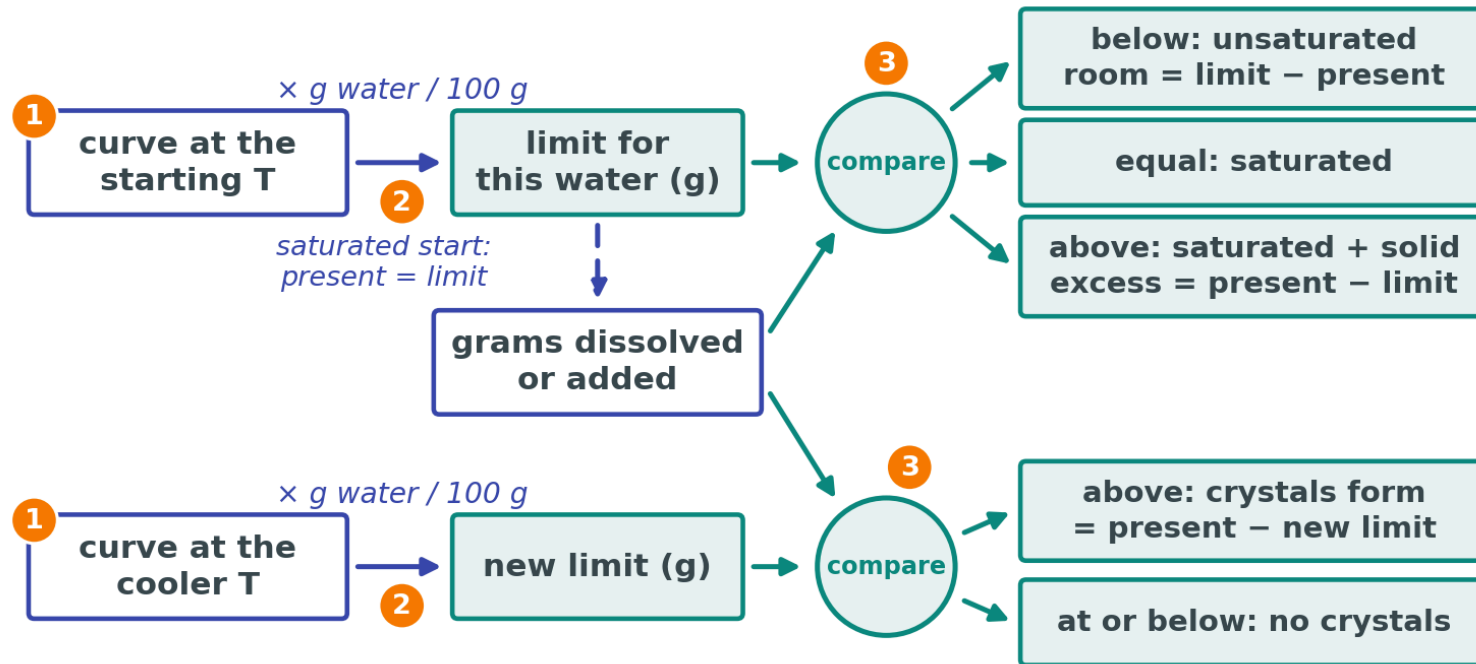
memory hook: hot tea takes more sugar · warm soda goes flat
most solids: more soluble when hot · gases: less soluble when hot



The method

1. **Read the curve at the temperature.** Grams per 100 g of water.
2. **Scale to the water present.** $\times (\text{grams of water} \div 100)$.
3. **Compare with the limit.** Below unsaturated, at saturated, above saturated plus solid.
Cooling drops the excess out.

The route through a saturation problem



Read the curve, scale it to the water, compare with the grams present. A cooled solution runs the lower lane too: the new limit sets what crystallizes.



Guided example: more solid than the water holds

50. g KNO_3 stirred into 100. g of water at 30 °C

given: 50. g KNO_3 · 100. g water · curve at 30 °C: 46 g per 100 g water · wanted: grams left undissolved, and the type of solution

A student stirs until nothing more changes. A common first attempt: 50. g is more than the 46 g limit, so the solution is supersaturated.

Find the grams left on the bottom. Then classify the solution.

✓ Guided example: solution

50. g KNO_3 in 100. g of water at 30 °C

given: 50. g KNO_3 · 100. g water · curve at 30 °C: 46 g per 100 g water

Step 1 · Read the curve at the temperature

Step 2 · Scale to the water present

$$100. \cancel{\text{g H}_2\text{O}} \times \frac{46 \text{ g KNO}_3}{100 \cancel{\text{g H}_2\text{O}}} = 46 \text{ g KNO}_3 \text{ can dissolve}$$

Step 3 · Compare with the limit

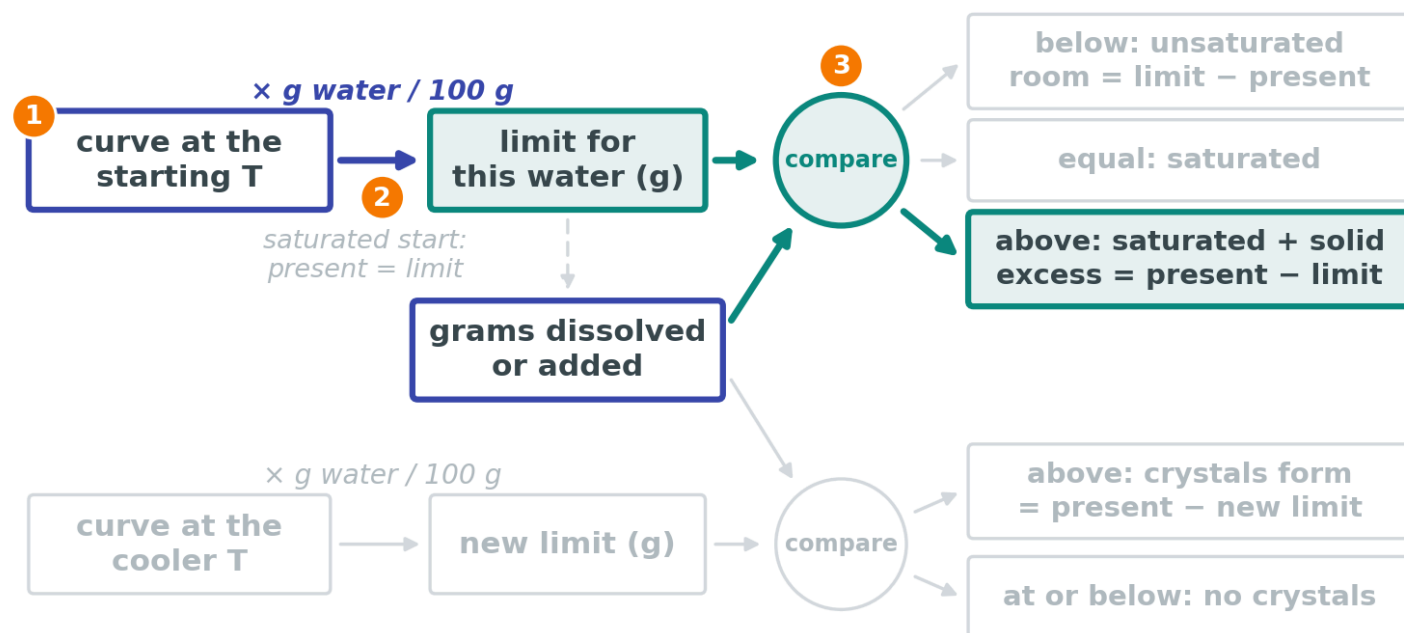
$$50. \text{ g added} - 46 \text{ g limit} = 4 \text{ g stay solid}$$

 **Sense check:** Solid resting in the liquid marks a saturated solution. A supersaturated one holds the extra dissolved, with no solid present. ✓

✓ Guided example: the route on the map

50. g KNO_3 in 100. g of water at 30 °C

given: 50. g KNO_3 · 100. g water · found: 46 g dissolve · 4 g stay solid · saturated



🧠 **Sense check:** Top lane: read, scale, compare. Above the limit, the extra stays solid and the solution is saturated. ✓



Worked example 1: classifying a solution

100 g KNO_3 stirred into 200 g of water at 40 °C: it all dissolves

given: 100 g solute · 200 g water · 40 °C · wanted: unsaturated, saturated, or supersaturated?

A common first attempt: the curve reads 64 g at 40 °C, so 64 g is the most this beaker holds. Test it.

Classify the solution.

✓ Worked example 1: solution

100 g KNO₃ in 200 g of water at 40 °C

given: 100 g solute · 200 g water · 40 °C

A common first attempt

Step 1 · Read the curve at the temperature

Step 2 · Scale to the water present

$$200 \text{ g H}_2\text{O} \times \frac{64 \text{ g KNO}_3}{100 \text{ g H}_2\text{O}} = 128 \text{ g KNO}_3$$

Step 3 · Compare with the limit

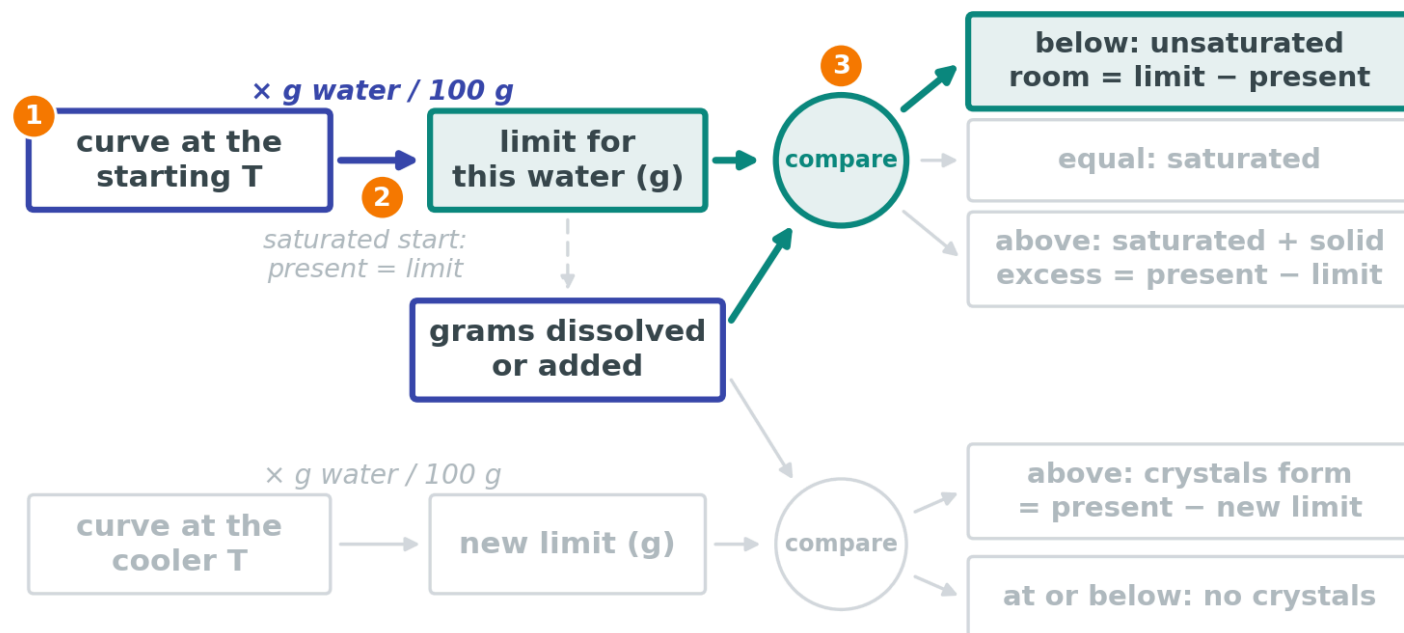
128 g limit – 100 g present = **28 g of room left → unsaturated**

 **Sense check:** The water could still take 28 g more KNO₃ before any solid appears. Below the limit means unsaturated.

✓ Worked example 1: the route on the map

100 g KNO_3 in 200 g of water at 40 °C

given: 100 g solute · 200 g water · 40 °C · found: limit 128 g · 28 g of room · unsaturated



🧠 **Sense check:** Top lane only. Scaling doubles the 64 g reading, and 100 g sits below the 128 g limit. ✓

Take-home: the curve value is per 100 g of water

curve read: 64 g per 100 g water at 40 °C
a ratio: the limit for this beaker depends on how much water it holds

200 g water $\rightarrow 200 \text{ g} \times (64 \text{ g} / 100 \text{ g}) = 128 \text{ g}$
scale before comparing: the beaker's limit is 128 g, not 64 g

Solubility is grams per 100 g of water. Multiply by the water actually present before judging saturation. Skipping the scale answers a different beaker.



Practice 1: the limit for this water

KCl at 20 °C: 34.0 g per 100 mL of water
a beaker holds 450. mL of water at 20 °C

What is the most KCl, in grams, that this water can dissolve?

- A. 153
- B. 7.56
- C. 34.0
- D. 1.32×10^3

✓ Practice 1 answer: A

KCl at 20 °C: 34.0 g per 100 mL water · 450. mL of water
the reading is per 100 mL, so scale it by the mL of water

$$450. \cancel{\text{ mL H}_2\text{O}} \times \frac{34.0 \text{ g KCl}}{100 \cancel{\text{ mL H}_2\text{O}}} = 153 \text{ g KCl} \cdot \text{answer A}$$

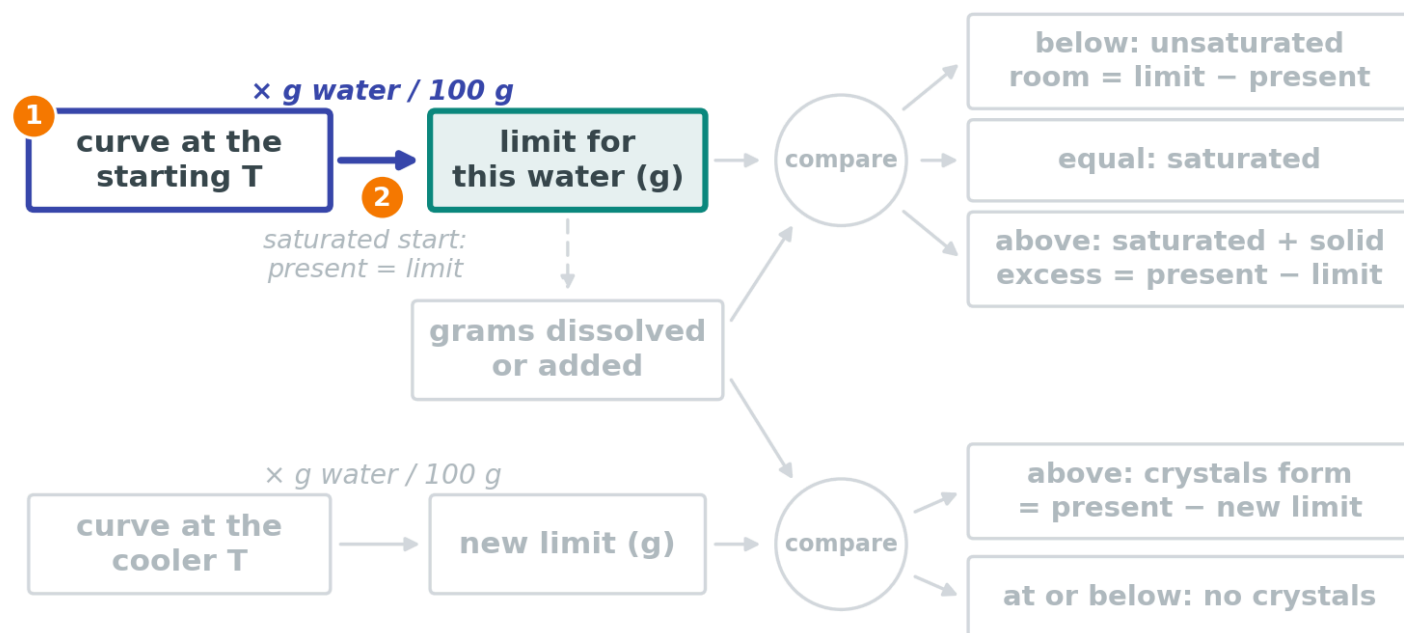
B scaled the wrong way: $34.0 \times (100/450.) = 7.56$. C used 34.0 g unscaled, the limit for 100 mL of water. D flipped the solubility: $450. \times (100/34.0) = 1.32 \times 10^3$, and mL H₂O does not cancel.

🧠 **Sense check:** 450. mL is 4.50 times 100 mL, so the limit is $4.50 \times 34.0 = 153 \text{ g}$. ✓

✓ Practice 1: the route on the map

KCl at 20 °C: 34.0 g per 100 mL water

given: 450. mL of water · found: 153 g KCl can dissolve



🧠 **Sense check:** Steps 1 and 2 only, scaled by mL of water. With no KCl added, there is nothing to compare. ✓



Practice 2: room left in the beaker

75.0 g of KI already dissolved in 125 g of water at 20 °C

KI at 20 °C: 144 g per 100 g of water

How many more grams of KI will this beaker take up?

- A. 180.
- B. 105
- C. 255
- D. 69.0

✓ Practice 2 answer: B

KI at 20 °C: 144 g per 100 g water · 75.0 g dissolved in 125 g of water
scale the limit to 125 g of water, then compare

$$125 \text{ g H}_2\text{O} \times \frac{144 \text{ g KI}}{100 \text{ g H}_2\text{O}} = 180. \text{ g limit}$$

$$180. \text{ g limit} - 75.0 \text{ g present} = \mathbf{105 \text{ g more} \cdot \text{answer B}}$$

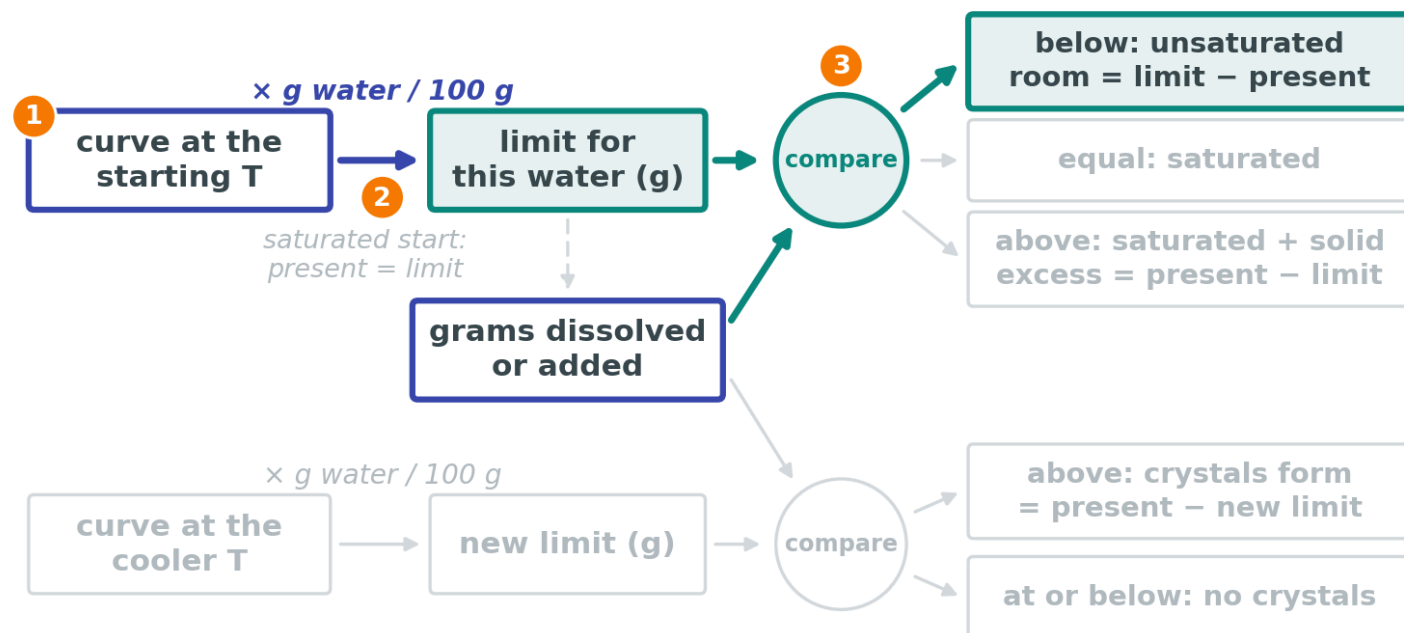
A stopped at the limit: 180. g counts the 75.0 g already in. C added instead of subtracting: $180. + 75.0 = 255$. D used 144 g unscaled: $144 - 75.0 = 69.0$.

🧠 **Sense check:** $75.0 \text{ g} + 105 \text{ g} = 180. \text{ g}$, the full limit for 125 g of water. ✓

✓ Practice 2: the route on the map

KI at 20 °C: 144 g per 100 g water

given: 75.0 g KI · 125 g water · found: limit 180. g · 105 g more · unsaturated



🧠 **Sense check:** The whole top lane: read, scale, compare. The grams present sit below the limit. ✓

Practice 3: three beakers of $\text{Ca}(\text{OH})_2$

$\text{Ca}(\text{OH})_2$ at 30 °C: 0.15 g per 100 g of water

beaker 1: 0.30 g in 200. g water · beaker 2: 0.50 g in 500. g water · beaker 3: 0.40 g in 120. g water

Each beaker is stirred at 30 °C until nothing more changes. Classify beakers 1, 2 and 3, in that order.

- A. saturated · unsaturated · supersaturated
- B. saturated · saturated · saturated
- C. saturated · unsaturated · saturated
- D. unsaturated · unsaturated · saturated

✓ Practice 3 answer: C

Ca(OH)_2 at 30 °C: 0.15 g per 100 g water

each limit = g of water $\times$ 0.15 g / 100 g water, then compare with the grams added

beaker	water	limit	added	solution
1	200. g	0.30 g	0.30 g	saturated
2	500. g	0.75 g	0.50 g	unsaturated
3	120. g	0.18 g	0.40 g	saturated + solid

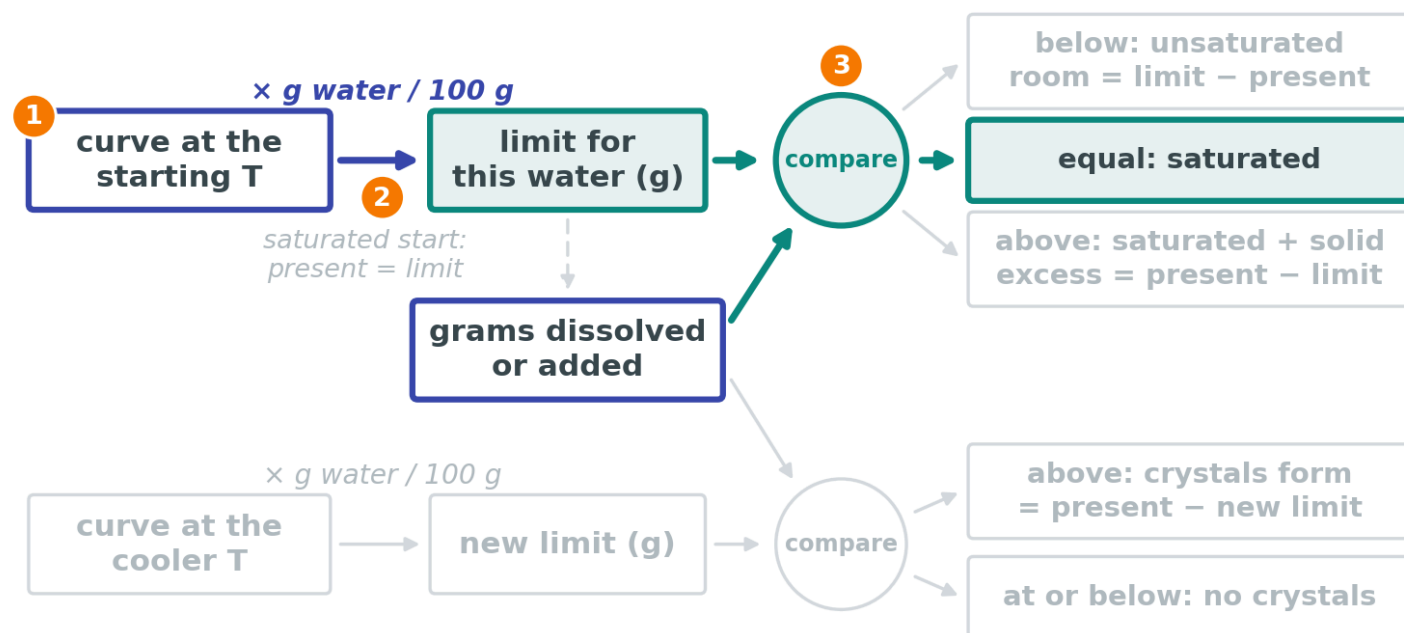
beaker 3: 0.40 g added – 0.18 g limit = **0.22 g stays solid** • answer C

A called beaker 3 supersaturated, but its extra stays solid. B compared every beaker with 0.15 g unscaled. D called beaker 1 unsaturated, yet 0.30 g is its full limit.

 **Sense check:** Beaker 2 holds the most Ca(OH)_2 yet has 0.25 g of room. The water sets the limit. ✓

✓ Practice 3: the route on the map

beaker 1: 0.30 g $\text{Ca}(\text{OH})_2$ in 200. g water at 30 °C
found: limit 0.30 g · 0.30 g added · saturated, no solid left



🧠 **Sense check:** Added equals the limit: saturated. Beakers 2 and 3 take the below and above boxes. ✓

Worked example 2: cooling a saturated solution

100 g of water holds all the KNO_3 it can at $60\text{ }^\circ\text{C}$, then cools to $20\text{ }^\circ\text{C}$
given: 100 g water · saturated at $60\text{ }^\circ\text{C}$ · cooled to $20\text{ }^\circ\text{C}$ · wanted: grams that crystallize

A hot solution saturated with KNO_3 is left to cool on the bench.

How many grams of KNO_3 fall out as crystals?

✓ Worked example 2: solution

saturated KNO_3 in 100 g water · $60\text{ }^\circ\text{C} \rightarrow 20\text{ }^\circ\text{C}$
 given: 100 g water · saturated at $60\text{ }^\circ\text{C}$ · cooled to $20\text{ }^\circ\text{C}$

Step 1 · Read the curve at the temperature

at $60\text{ }^\circ\text{C} = 106\text{ g} / 100\text{ g water}$ · at $20\text{ }^\circ\text{C} = 32\text{ g} / 100\text{ g water}$

Step 2 · Scale to the water present

Step 3 · Compare with the limit

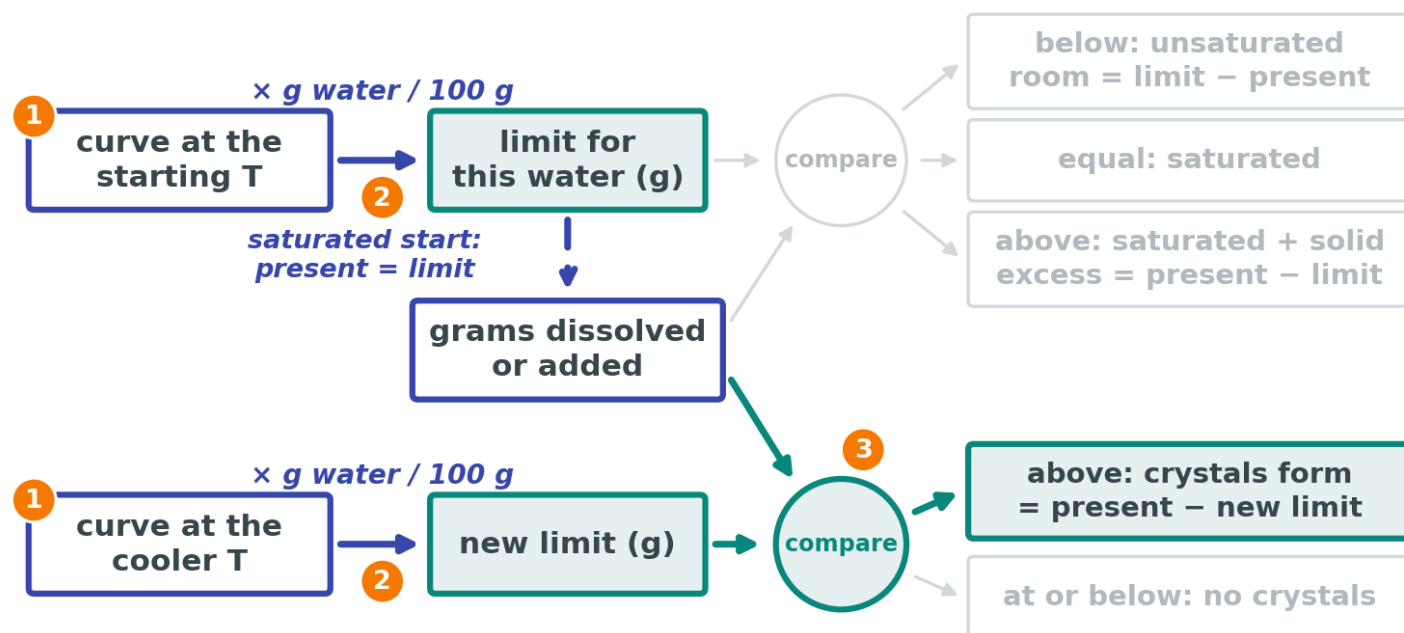
$106\text{ g dissolved} - 32\text{ g the cool water can hold} = 74\text{ g crystallize}$

 **Sense check:** Cooling lowers the limit, and the 74 g it can no longer hold drops out as crystals. Less dissolves cold than hot.

✓ Worked example 2: the route on the map

saturated KNO_3 in 100 g water · $60^\circ\text{C} \rightarrow 20^\circ\text{C}$

given: 100 g water · saturated at 60°C · cooled to 20°C · found: 74 g crystallize



🧠 **Sense check:** Saturated at the start, so the 60°C limit is the grams dissolved. The 20°C lane sets the new limit. ✓

Your turn: cooling from 50 °C to 30 °C

saturated KNO_3 in 100 g of water at 50 °C, cooled to 30 °C

curve: 50 °C reads 85 g / 100 g water · 30 °C reads 46 g / 100 g water

85 g dissolved – 46 g the cool water can hold = _____ g crystallize

Read both limits, then subtract to find what falls out.

85 g dissolved – 46 g the cool water can hold = **39 g crystallize**

Cooling from 50 °C to 30 °C drops 39 g of KNO_3 out of solution.

⚠ Where this goes wrong

Stirring past the limit. Add 90 g of a salt whose limit is 60 g, then stir for an hour: 60 g dissolves and $90 - 60 = 30$ g stays on the bottom. Stirring speeds dissolving up to the limit; it cannot push past it.

Using the curve value without scaling. The curve gives grams per 100 g of water. For 150 g of water at 60 °C the beaker holds $150 \times (106 \text{ g} / 100 \text{ g}) = 159$ g, not 106 g. Scale to the water present first.

Adding instead of subtracting on cooling. A solution with 106 g dissolved, cooled to a 32 g limit, does not shed $106 + 32 = 138$ g. What crystallizes is the excess: $106 - 32 = 74$ g.

Reading at the wrong temperature. Cooling to 20 °C means reading the limit at 20 °C, not at 40 °C. The 40 °C value gives $106 - 64 = 42$ g and understates the crystals. Read straight up from the final temperature.



Practice 4

NaCl solubility at 20 °C: 36 g per 100 g of water
75 g of NaCl stirred into 50. g of water at 20 °C, stirred thoroughly

How many grams of NaCl remain undissolved at the bottom?

- A. 0
- B. 18
- C. 39
- D. 57

✓ Practice 4 answer: D

NaCl at 20 °C: 36 g per 100 g water · 75 g into 50. g water
 scale the limit to 50. g of water, then subtract

$$50. \text{ g H}_2\text{O} \times \frac{36 \text{ g NaCl}}{100 \text{ g H}_2\text{O}} = 18 \text{ g dissolve}$$

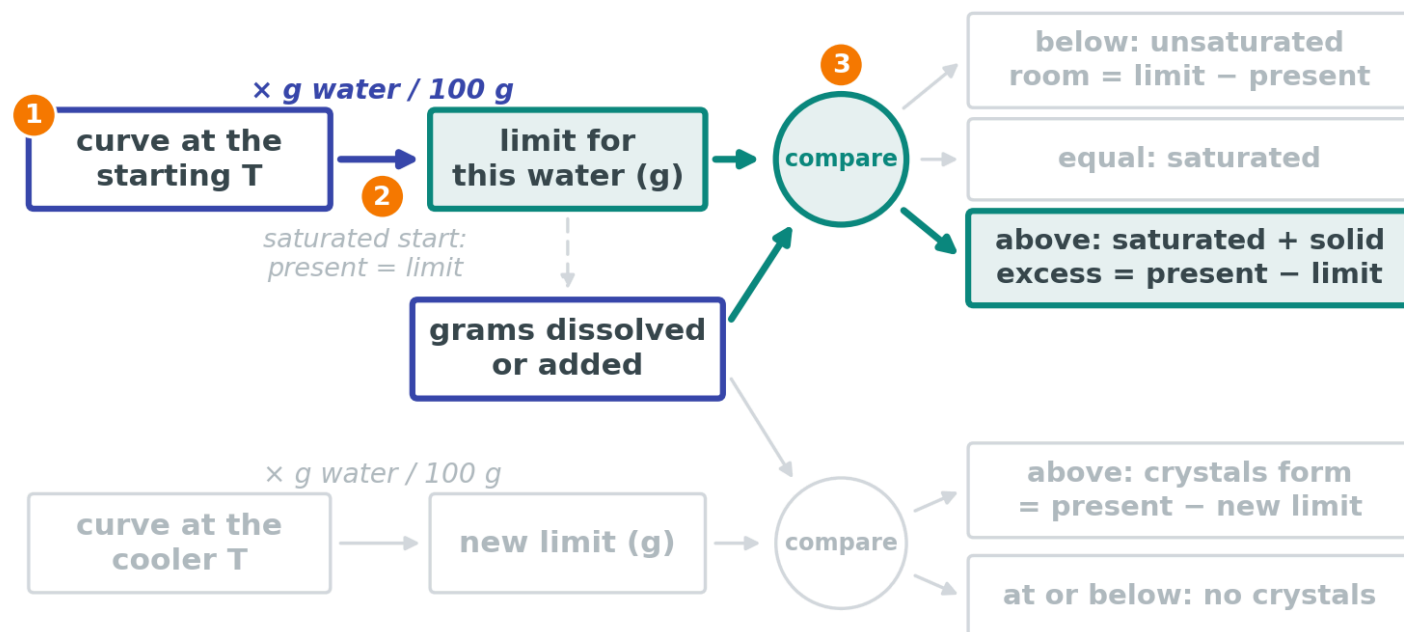
75 g added – 18 g dissolved = **57 g undissolved · answer D**

A: 0 g assumes stirring dissolves everything; past the limit the extra stays solid. B: 18 g is how much dissolves, not what is left: $75 - 18 = 57$. C: 39 g used 36 g as the limit for this beaker, but 36 g is per 100 g of water: only 50. g is here, so $50. \times 36/100 = 18$ g dissolves.

 **Sense check:** More salt was added than 50. g of water can hold, so solid must remain. The 18 g that dissolves leaves 57 g on the bottom. ✓

✓ Practice 4: the route on the map

NaCl at 20 °C: 36 g per 100 g water · 75 g into 50. g water
found: 18 g dissolve · 57 g stay solid



🧠 **Sense check:** Scaling halves the 36 g reading. 75 g is far above the 18 g limit, so the excess stays on the bottom. ✓

Practice 5: cooling a saturated solution

$\text{K}_2\text{Cr}_2\text{O}_7$: 73.0 g per 100 g water at 80 °C · 12.0 g per 100 g water at 20 °C
a solution saturated with $\text{K}_2\text{Cr}_2\text{O}_7$ in 40.0 g of water at 80 °C is cooled to 20 °C

How many grams of $\text{K}_2\text{Cr}_2\text{O}_7$ crystallize?

- A. 24.4
- B. 4.80
- C. 61.0
- D. 17.2

✓ Practice 5 answer: A

saturated $\text{K}_2\text{Cr}_2\text{O}_7$ in 40.0 g water · 80 °C → 20 °C
scale both limits to 40.0 g of water

$$40.0 \cancel{\text{ g H}_2\text{O}} \times \frac{73.0 \text{ g K}_2\text{Cr}_2\text{O}_7}{100 \cancel{\text{ g H}_2\text{O}}} = 29.2 \text{ g dissolved at } 80 \text{ °C}$$

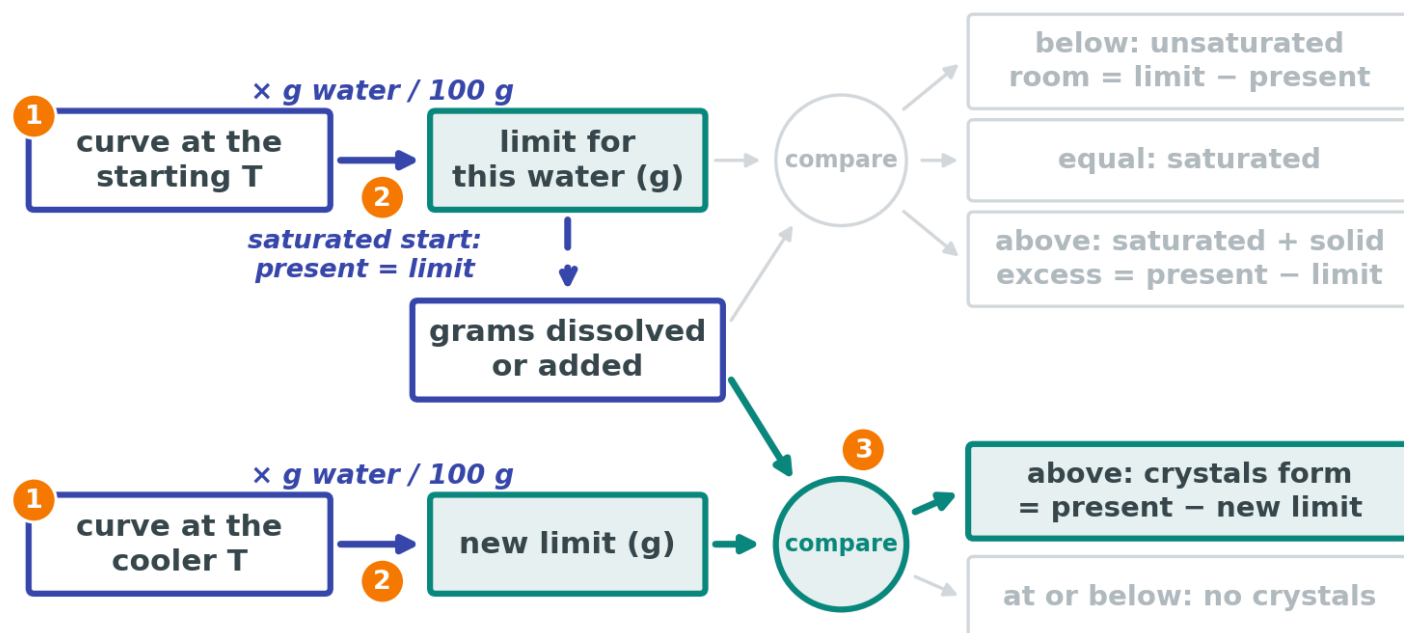
$$40.0 \cancel{\text{ g H}_2\text{O}} \times \frac{12.0 \text{ g K}_2\text{Cr}_2\text{O}_7}{100 \cancel{\text{ g H}_2\text{O}}} = 4.80 \text{ g stays} \cdot 29.2 - 4.80 = \mathbf{24.4 \text{ g crystallize} \cdot \text{answer A}}$$

B stopped at 4.80 g, what stays dissolved. C used both readings unscaled: $73.0 - 12.0 = 61.0$. D scaled the 80 °C reading but not the 20 °C one: $29.2 - 12.0 = 17.2$.

🧠 **Sense check:** 40.0 g of water is 0.400 of 100 g, so $0.400 \times 61.0 = 24.4 \text{ g crystallize}$. ✓

✓ Practice 5: the route on the map

saturated $\text{K}_2\text{Cr}_2\text{O}_7$ in 40.0 g water · $80\text{ }^\circ\text{C} \rightarrow 20\text{ }^\circ\text{C}$
 found: 29.2 g dissolved · new limit 4.80 g · 24.4 g crystallize



🧠 **Sense check:** Saturated start: the $80\text{ }^\circ\text{C}$ limit is the grams dissolved. Both lanes scale by 40.0 g / 100 g. ✓

 Practice 6

KNO₃ curve: 136 g / 100 g water at 70 °C · 46.0 g / 100 g water at 30 °C

180. g of KNO₃ is dissolved in 250. g of water at 70 °C, then cooled to 30 °C

How many grams of KNO₃ crystallize?

- A. 225
- B. 115
- C. 65
- D. 134

✓ Practice 6 answer: C

180. g KNO₃ in 250. g water · cooled from 70 °C to 30 °C
scale both limits to 250. g of water

$$250. \cancel{\text{g H}_2\text{O}} \times \frac{136 \text{ g KNO}_3}{100 \cancel{\text{g H}_2\text{O}}} = 340. \text{ g limit at } 70 \text{ °C} \cdot 180. \text{ g present: } \mathbf{\text{unsaturated}}$$

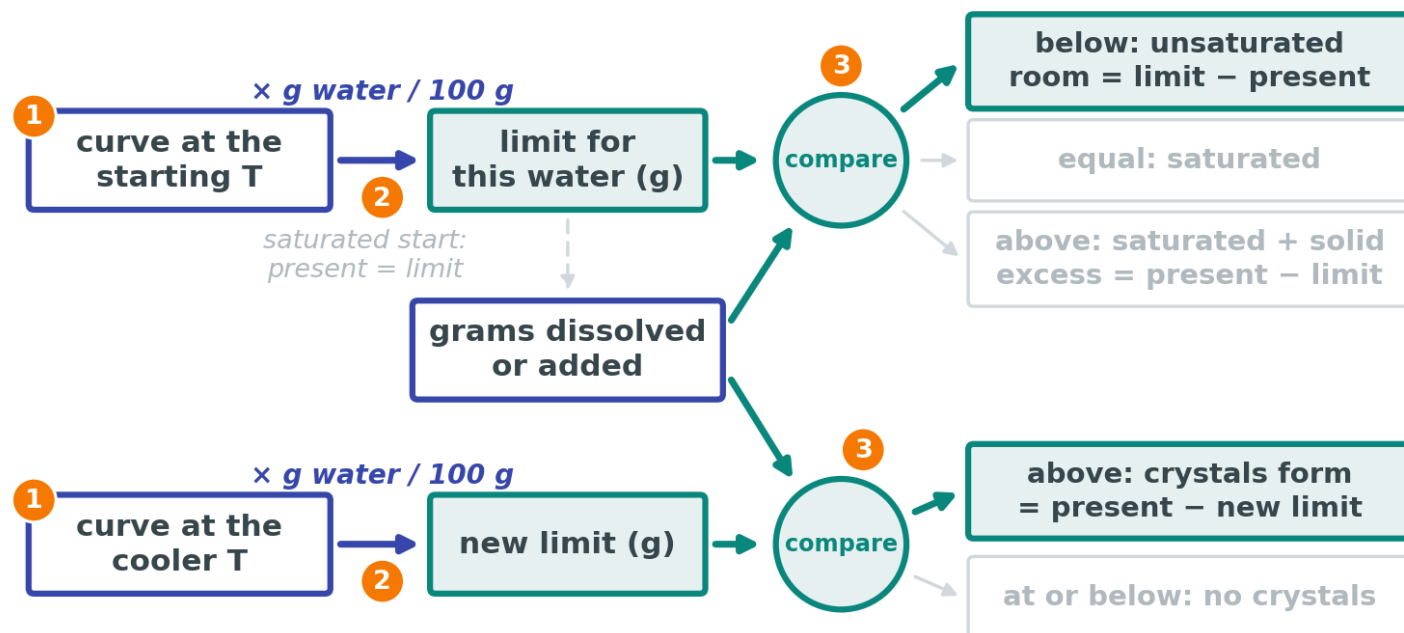
$$250. \cancel{\text{g H}_2\text{O}} \times \frac{46.0 \text{ g KNO}_3}{100 \cancel{\text{g H}_2\text{O}}} = 115 \text{ g stays} \cdot 180. - 115 = \mathbf{65 \text{ g crystallize} \cdot \text{answer C}}$$

A assumed a saturated start: 340. – 115 = 225 g, but only 180. g was dissolved. B stopped at 115 g, what stays dissolved. D used 46.0 g unscaled: 180. – 46.0 = 134 g.

 **Sense check:** Only what was dissolved can crystallize: 180. g in, 115 g stays, 65 g out. ✓

✓ Practice 6: the route on the map

180. g KNO_3 in 250. g water · 70 °C → 30 °C
 found: 340. g limit at 70 °C · 115 g limit at 30 °C · 65 g crystallize



Sense check: Both lanes: all 180. g dissolve at 70 °C, and the excess over 115 g crystallizes. ✓

 Practice 7

K_2SO_4 at 40 °C: 14.8 g per 100 g of water
25.0 g of K_2SO_4 must dissolve completely at 40 °C

What is the least mass of water, in grams, that dissolves all of it?

- A. 100.
- B. 3.70
- C. 10.2
- D. 169

✓ Practice 7 answer: D

K_2SO_4 at 40 °C: 14.8 g per 100 g water · 25.0 g K_2SO_4
the solubility is a conversion factor: g K_2SO_4 goes below so it cancels

$$25.0 \cancel{\text{ g K}_2\text{SO}_4} \times \frac{100 \text{ g H}_2\text{O}}{14.8 \cancel{\text{ g K}_2\text{SO}_4}} = 169 \text{ g H}_2\text{O} \cdot \text{answer D}$$

A took the 100 g of water in the reading, but 100 g of water holds only 14.8 g. B flipped the factor: $25.0 \times 14.8/100 = 3.70$, and g K_2SO_4 does not cancel. C subtracted: $25.0 - 14.8 = 10.2$ g is what stays solid in 100 g of water.

🧠 **Sense check:** 25.0 g is 1.69 times 14.8 g, so it needs 1.69 times 100 g of water: 169 g ✓

 Practice 8

NaNO_3 at 20 °C: 88.0 g per 100 g of water
a flask holds 175. g of saturated NaNO_3 solution at 20 °C

How many grams of NaNO_3 are dissolved in the flask?

- A. 81.9
- B. 154
- C. 88.0
- D. 93.1

✓ Practice 8 answer: A

saturated NaNO_3 at 20 °C: 88.0 g NaNO_3 + 100 g water = 188.0 g solution

175. g of solution, not of water · the solution mass goes below

$$175. \text{ g } \cancel{\text{solution}} \times \frac{88.0 \text{ g } \text{NaNO}_3}{188.0 \text{ g } \cancel{\text{solution}}} = 81.9 \text{ g } \text{NaNO}_3 \cdot \text{answer A}$$

B divided by the water, not the solution: $175. \times 88.0/100 = 154$, as if the flask held 175. g of water. C used 88.0 g unscaled, the NaNO_3 in 188.0 g of solution. D gave the water: $175. \times 100/188.0 = 93.1$.

🧠 **Sense check:** 88.0 of every 188.0 g is NaNO_3 , 46.8% by mass: $0.468 \times 175. = 81.9 \text{ g } \checkmark$

 Practice 9

NH_4Cl : 65.6 g per 100 g of water at 80 °C · 37.2 g per 100 g of water at 20 °C

NH_4Cl saturates 125 g of water at 80 °C, then the beaker cools to 20 °C and crystals form

How many grams of water, added at 20 °C, redissolve the crystals?

- A. 54.1
- B. 95.4
- C. 35.5
- D. 13.2

✓ Practice 9 answer: B

saturated NH_4Cl in 125 g water · $80\text{ }^\circ\text{C} \rightarrow 20\text{ }^\circ\text{C}$

65.6 g per 100 g water at $80\text{ }^\circ\text{C}$ · 37.2 g at $20\text{ }^\circ\text{C}$ · the crystals first, then the water that holds them

$$125 \text{ g } \cancel{\text{H}_2\text{O}} \times \frac{(65.6 - 37.2) \text{ g } \text{NH}_4\text{Cl}}{100 \text{ g } \cancel{\text{H}_2\text{O}}} = 35.5 \text{ g } \text{NH}_4\text{Cl crystallize}$$

$$35.5 \text{ g } \cancel{\text{NH}_4\text{Cl}} \times \frac{100 \text{ g } \text{H}_2\text{O}}{37.2 \text{ g } \cancel{\text{NH}_4\text{Cl}}} = 95.4 \text{ g } \text{H}_2\text{O} \cdot \text{answer B}$$

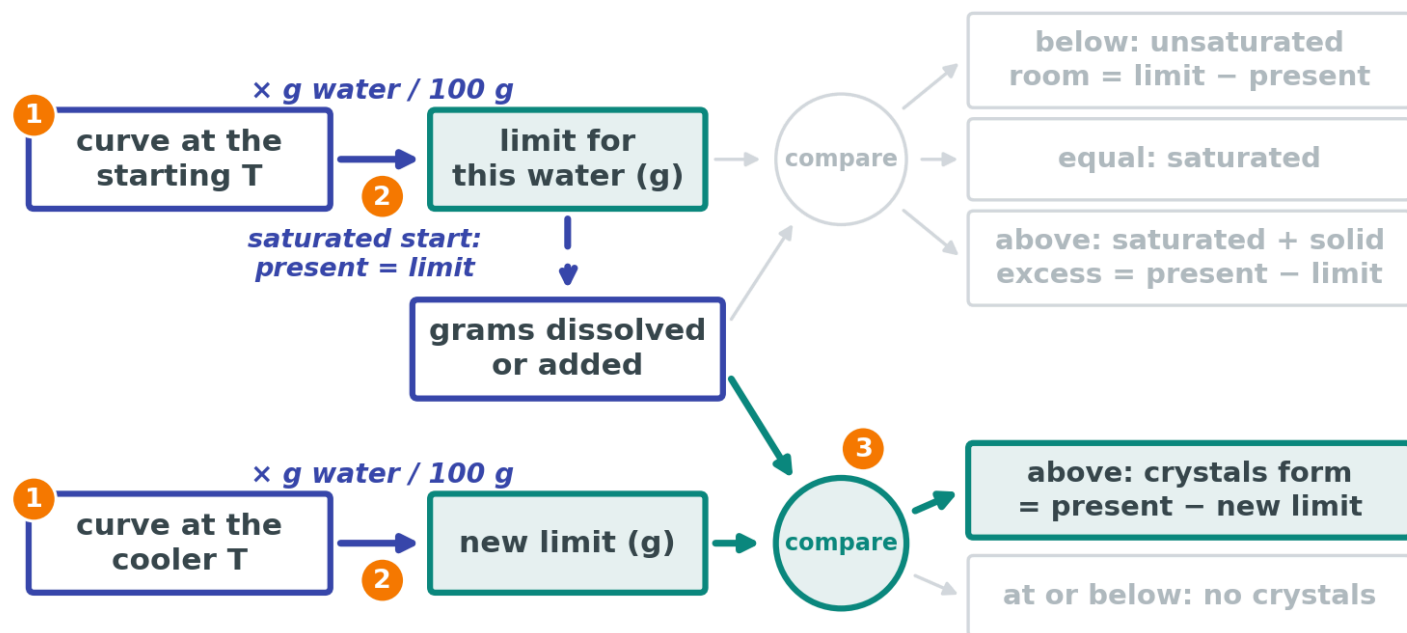
A used the $80\text{ }^\circ\text{C}$ reading: $35.5 \times 100/65.6 = 54.1$, but the water goes in at $20\text{ }^\circ\text{C}$. C stopped at the crystals, 35.5 g. D flipped the factor: $35.5 \times 37.2/100 = 13.2$.

 **Sense check:** 35.5 g is a little under 37.2 g, so it needs a little under 100 g of water: 95.4 g ✓

✓ Practice 9: the route on the map

saturated NH_4Cl in 125 g water · $80\text{ }^\circ\text{C} \rightarrow 20\text{ }^\circ\text{C}$

found: 82.0 g dissolved · new limit 46.5 g · 35.5 g crystallize · 95.4 g water redissolves them



🧠 **Sense check:** The map ends at 35.5 g of crystals. Then flip the $20\text{ }^\circ\text{C}$ reading, as in Practice 7. ✓

Check yourself

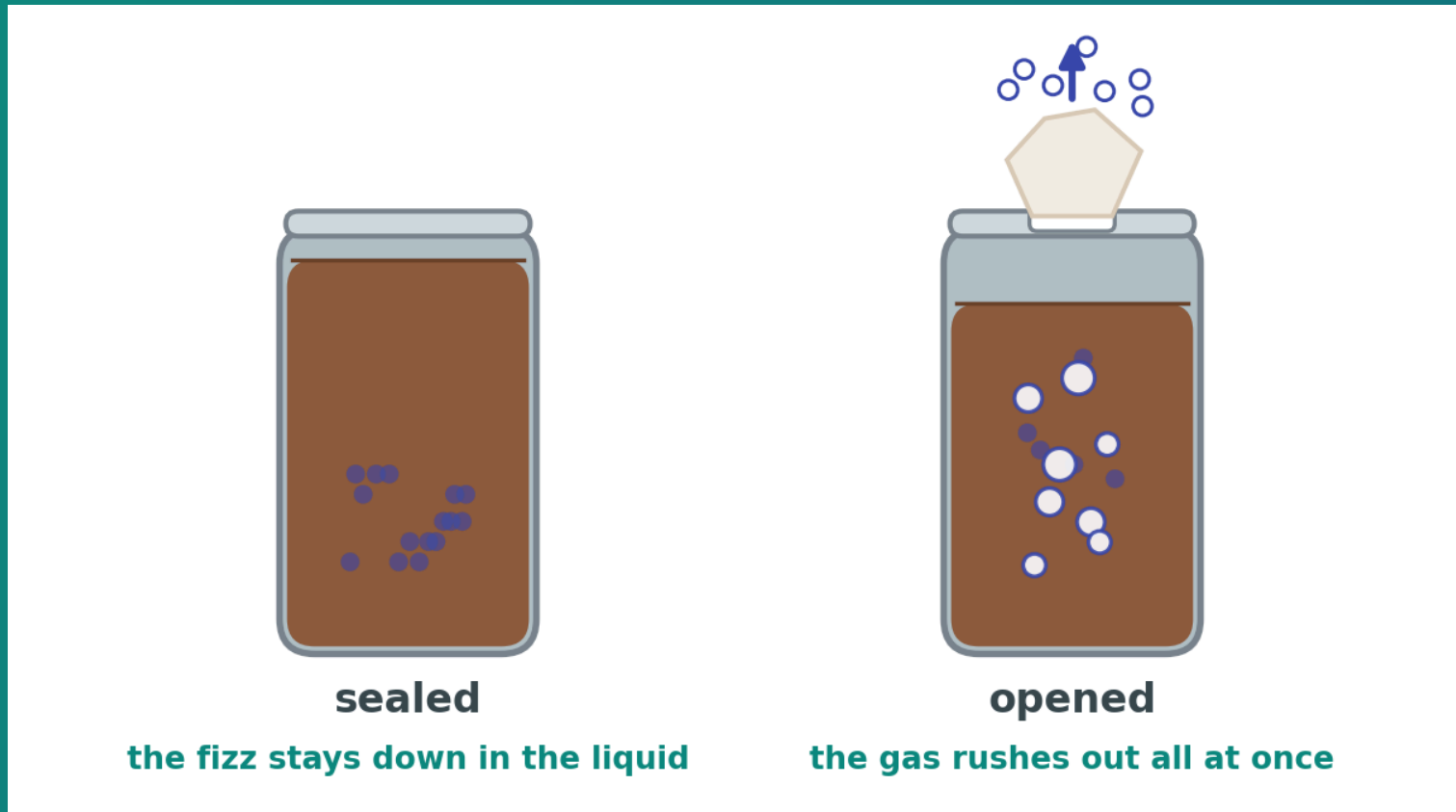
1. A solution holds 30 g of KNO_3 in 100 g of water at 60 °C, where the limit is 106 g. Unsaturated, saturated, or supersaturated?
2. A solution saturated with KNO_3 at 40 °C (limit 64 g) in 100 g of water is cooled to 20 °C (limit 32 g). How many grams crystallize?

A dissolved solute raises the next question: how much of it is in the water: the **concentration**. Adding water lowers the concentration without changing the amount of solute, the relation $M_1V_1 = M_2V_2$.

5 • Henry's Law

Use Henry's law (a gas's solubility is directly proportional to its partial pressure, $S_1/P_1 = S_2/P_2$) to find a new solubility, or the pressure that produces it, at constant temperature.

Open a soda and it erupts



Crack the tab on a shaken can and it hisses, then foams over. Sealed, the fizz stayed down in the drink. Opened, the gas escapes all at once.

Pressure changes gas solubility only

three factors set solubility: structure · temperature · pressure
pressure matters for a gas solute only · a dissolved solid or liquid does not respond

Pressure pushes a gas into a liquid. A dissolved solid does not respond.

Sugar water and soda water are sealed under more pressure. Which solubility rises?

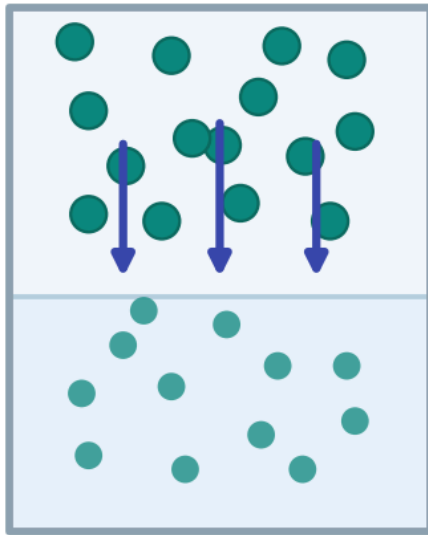
 **Sense check:** The CO₂. Sugar is a solid. ✓

More gas overhead, more gas dissolved

more gas pressing on the surface → more gas dissolved

more gas overhead

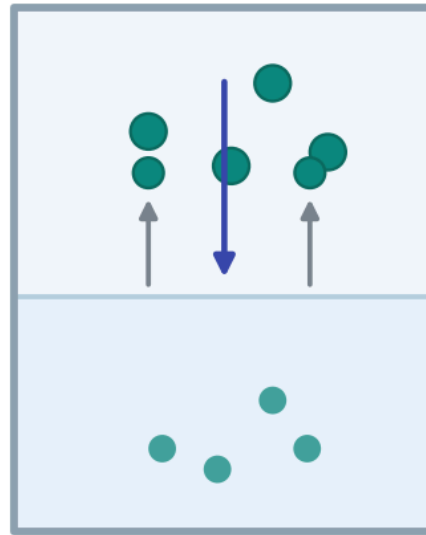
high pressure



much dissolves

less gas overhead

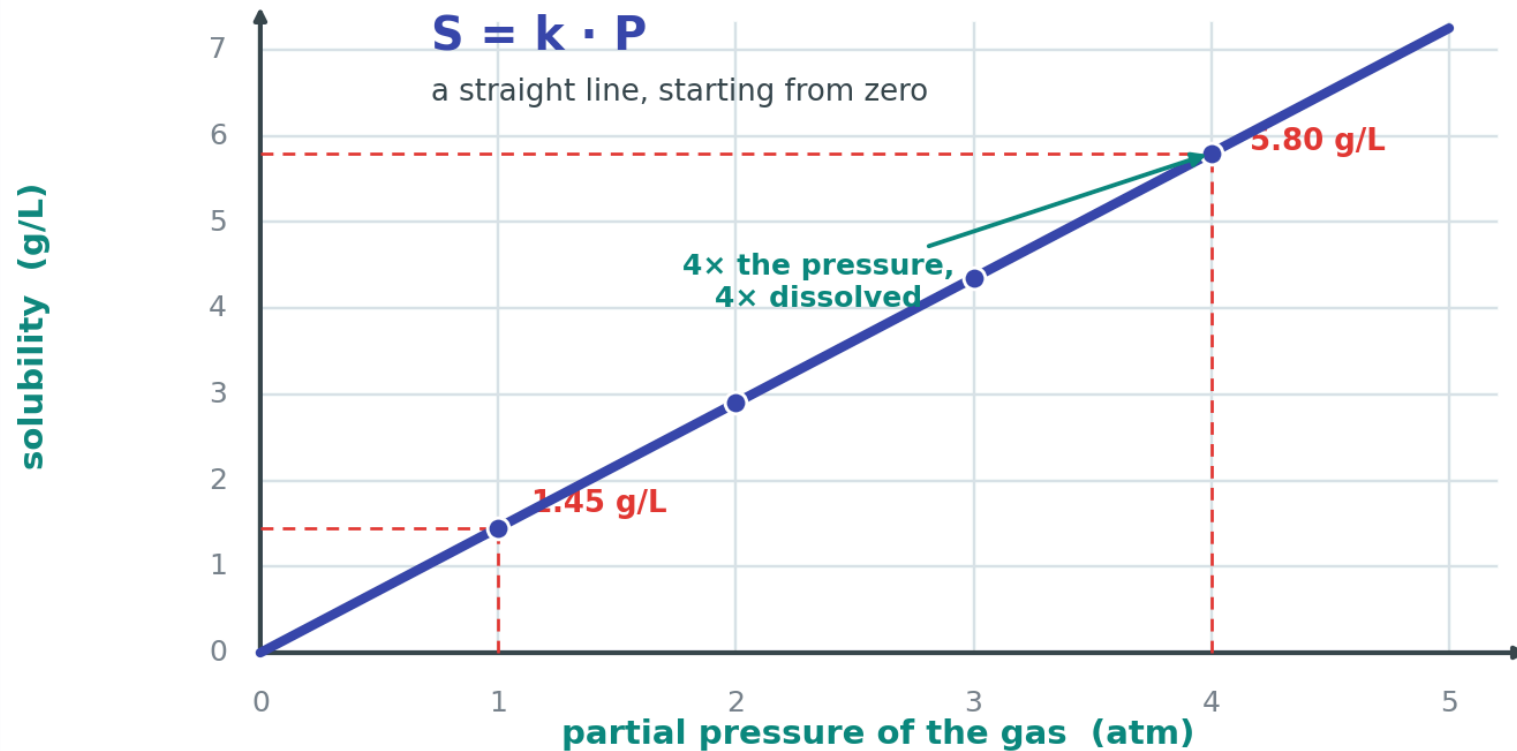
low pressure



little dissolves

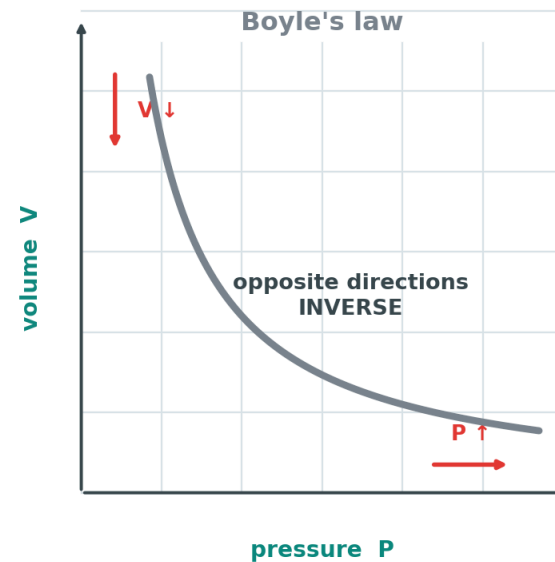
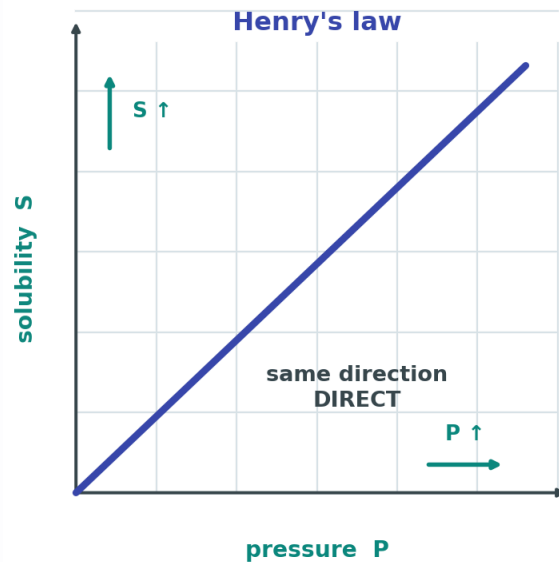
At constant temperature, a gas's solubility (how much dissolves) rises and falls in step with its partial pressure. More gas overhead drives more into solution.

Solubility is proportional to pressure



Plot the dissolved amount against the partial pressure: the points fall on a straight line through the origin. Double the pressure, double the dissolved gas. The slope depends only on the gas and temperature.

Henry's law is direct, not inverse



Raise the pressure and solubility rises with it. The two move the same way. Boyle's law is the opposite: squeeze a trapped gas and its volume shrinks. Same variable, but opposite behavior.

memory hook: pressure up, gas in · pressure down, gas out
a sealed soda holds its fizz · an opened one goes flat

Henry's law in symbols

$$C_1/P_1 = C_2/P_2, \text{ also written } S_1/P_1 = S_2/P_2$$

C or S: the gas's solubility, a concentration · P: its partial pressure · same gas, constant temperature

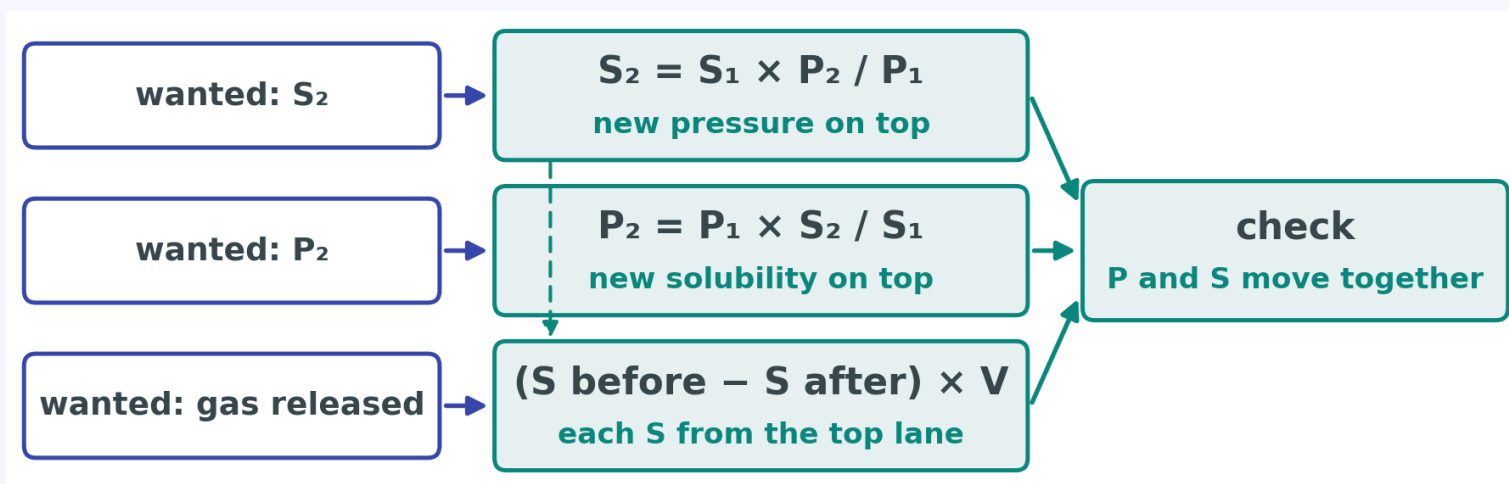
Any concentration unit works for S: g/L, mg/L, g/100 mL or M. Any pressure unit works for P: atm or mmHg. Keep each pair in one unit so it cancels.

$$S_2 = S_1 \text{ in mg/L} \times \frac{P_2 \text{ in mmHg}}{P_1 \text{ in mmHg}} = S_2 \text{ in mg/L}$$



The method

1. **Identify given and wanted.** Mark the unknown.
2. **Set up the proportion.** Same gas, constant temperature: $S_1/P_1 = S_2/P_2$.
3. **Rearrange for the unknown.**
4. **Substitute and check.** More pressure means more dissolved gas.



Guided example: a mountain lake

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 9.0 \text{ mg/L}$ · $P_1 = 160 \text{ mmHg}$ · $P_2 = 80. \text{ mmHg}$ · same gas, constant temperature · wanted: S_2

A lake at sea level holds **9.0 mg/L** of dissolved O_2 under **160 mmHg** of O_2 . A high mountain lake at the same temperature sits under only **80. mmHg** of O_2 . How much O_2 can it hold, in mg/L?

Compare the pressures first: 80. mmHg is half of 160 mmHg.

✓ Guided example: solution

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 9.0 \text{ mg/L}$ · $P_1 = 160 \text{ mmHg}$ · $P_2 = 80. \text{ mmHg}$ · wanted: S_2

Step 1 · Identify given and wanted

Step 2 · Set up the proportion

Step 3 · Rearrange for the unknown

$$S_1/P_1 = S_2/P_2 \rightarrow S_2 = S_1 \times \frac{P_2}{P_1}$$

Step 4 · Substitute and check

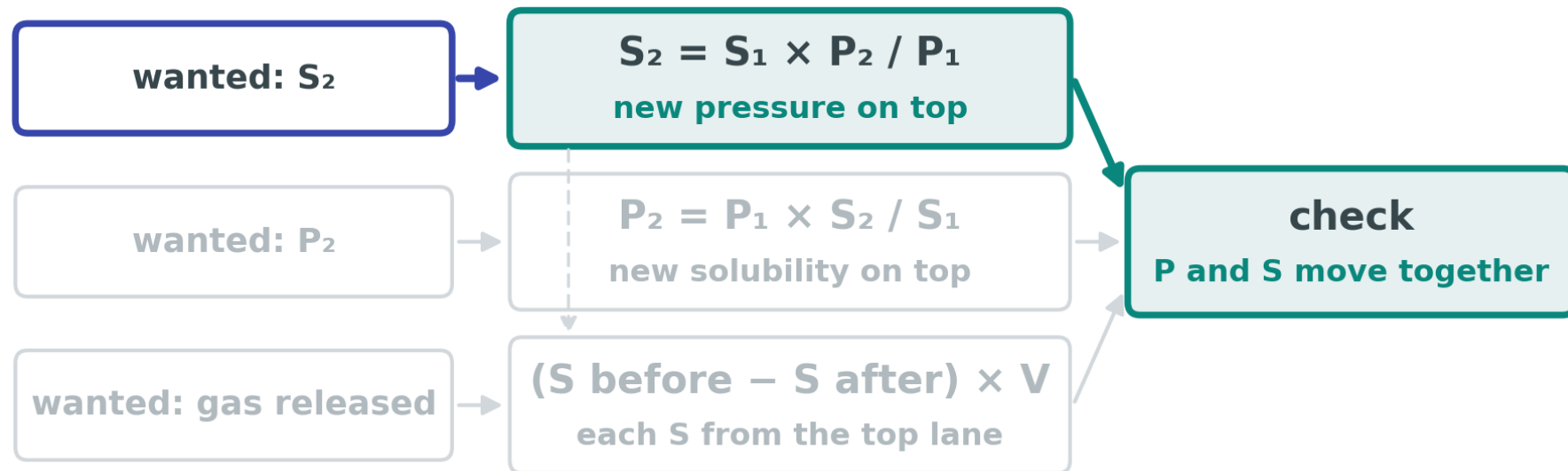
$$S_2 = 9.0 \text{ mg/L} \times \frac{80. \cancel{\text{ mmHg}}}{160 \cancel{\text{ mmHg}}} = 4.5 \text{ mg/L}$$

 **Sense check:** The O_2 pressure halved, so half as much O_2 dissolves: $9.0 \rightarrow 4.5 \text{ mg/L}$. mmHg cancels and mg/L carries through. ✓

✓ Guided example: the route on the strip

$$S_1/P_1 = S_2/P_2$$

given: 9.0 mg/L at 160 mmHg · $P_2 = 80.$ mmHg · found: $S_2 = 4.5$ mg/L



 **Sense check:** A solubility is wanted, so the top lane: the new pressure goes on top. The pressure fell, and the solubility fell with it. ✓



Practice 1

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 42.0 \text{ mg/L}$ · $P_1 = 2.40 \text{ atm}$ · $P_2 = 0.800 \text{ atm}$ · same gas, constant temperature · wanted: S_2

At depth, a diver's blood holds **42.0 mg/L** of N_2 under **2.40 atm** of N_2 . At the surface the N_2 pressure is **0.800 atm**. How many milligrams of N_2 per liter can her blood hold there?

- A. 126
- B. 28.0
- C. 42.0
- D. 14.0

✓ Practice 1 answer: D

$$S_1/P_1 = S_2/P_2$$

$$N_2 \cdot S_1 = 42.0 \text{ mg/L at } P_1 = 2.40 \text{ atm} \cdot P_2 = 0.800 \text{ atm} \cdot \text{constant temperature}$$

$$S_2 = 42.0 \text{ mg/L} \times \frac{0.800 \text{ atm}}{2.40 \text{ atm}} = 14.0 \text{ mg/L} \cdot \text{answer D}$$

A inverted the proportion like Boyle's law: $42.0 \times (2.40/0.800) = 126 \text{ mg/L}$, more N_2 held at lower pressure. B scaled by the pressure change: $42.0 \times (2.40 - 0.800)/2.40 = 28.0 \text{ mg/L}$, the N_2 that comes out, not what stays. C assumed no change: at a third of the pressure, 42.0 mg/L cannot stay dissolved.

 **Sense check:** The pressure fell to a third, so a third of the N_2 stays dissolved: $42.0 \rightarrow 14.0 \text{ mg/L}$. The other 28.0 mg/L leaves the blood. ✓



Worked example 1: more pressure, more dissolved

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 1.45 \text{ g/L}$ · $P_1 = 1.0 \text{ atm}$ · $P_2 = 4.0 \text{ atm}$ · same gas, constant temperature · wanted: S_2

Water sits under carbon dioxide. At **1.0 atm** of CO_2 , **1.45 g/L** dissolves. The CO_2 pressure is raised to **4.0 atm** at constant temperature. Find the new solubility.

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

✓ Worked example 1: solution

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 1.45 \text{ g/L}$ · $P_1 = 1.0 \text{ atm}$ · $P_2 = 4.0 \text{ atm}$ · same gas, constant temperature · wanted: S_2

Step 1 · Identify given and wanted

Step 2 · Set up the proportion

Step 3 · Rearrange for the unknown

$$S_1/P_1 = S_2/P_2 \rightarrow S_2 = S_1 \times \frac{P_2}{P_1}$$

Step 4 · Substitute and check

$$S_2 = 1.45 \text{ g/L} \times \frac{4.0 \text{ atm}}{1.0 \text{ atm}} = 5.80 \text{ g/L}$$

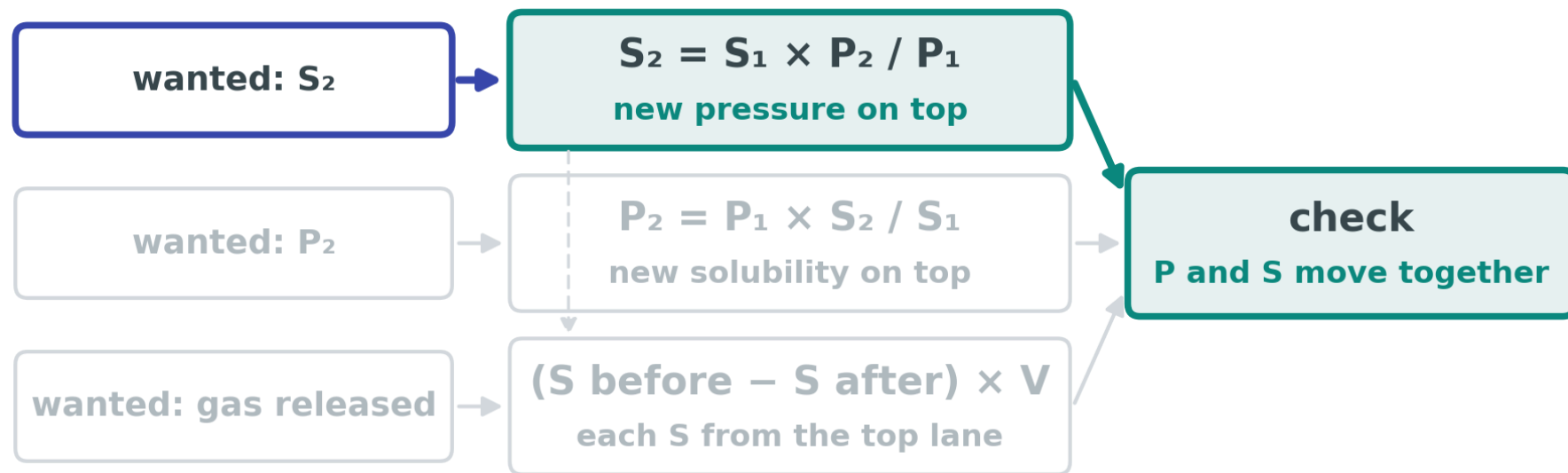
 **Sense check:** Pressure quadrupled at constant temperature, so four times as much CO_2 dissolves: $1.45 \rightarrow 5.80 \text{ g/L}$.

✓

✓ Worked example 1: the route on the strip

$$S_1/P_1 = S_2/P_2$$

given: 1.45 g/L at 1.0 atm · $P_2 = 4.0$ atm · found: $S_2 = 5.80$ g/L



🧠 **Sense check:** A solubility is wanted, so the top lane again. The pressure rose, and the solubility rose with it. ✓



Worked example 2: pressure released, gas comes out

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 5.80 \text{ g/L}$ · $P_1 = 4.0 \text{ atm}$ · $P_2 = 1.0 \text{ atm}$ · same gas, constant temperature · wanted: S_2

A carbonation tank holds water under **4.0 atm** of CO_2 , dissolving **5.80 g/L**. The CO_2 pressure is bled down to **1.0 atm**. Find the new dissolved amount.

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

✓ Worked example 2: solution

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 5.80 \text{ g/L}$ · $P_1 = 4.0 \text{ atm}$ · $P_2 = 1.0 \text{ atm}$ · same gas, constant temperature · wanted: S_2

Step 1 · Identify given and wanted

Step 2 · Set up the proportion

Step 3 · Rearrange for the unknown

$$S_1/P_1 = S_2/P_2 \rightarrow S_2 = S_1 \times \frac{P_2}{P_1}$$

Step 4 · Substitute and check

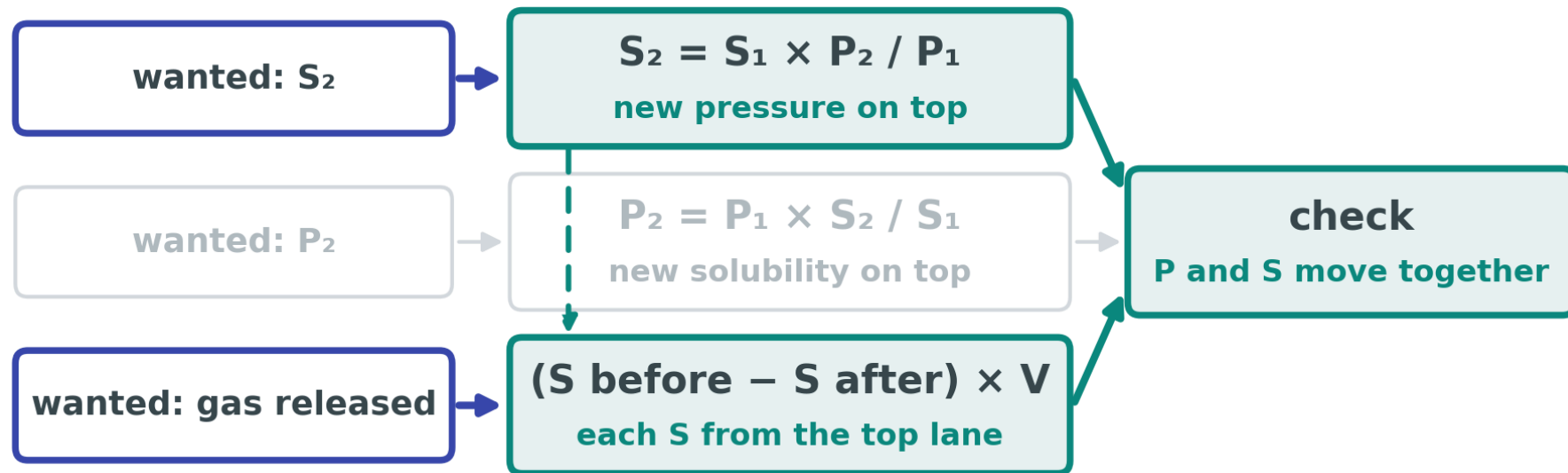
$$S_2 = 5.80 \text{ g/L} \times \frac{1.0 \text{ atm}}{4.0 \text{ atm}} = 1.45 \text{ g/L}$$

 **Sense check:** Pressure dropped to a quarter, so only a quarter of the CO_2 stays dissolved: $5.80 \rightarrow 1.45 \text{ g/L}$. The rest, $5.80 - 1.45 = 4.35 \text{ g}$ from every liter, fizzes out. ✓

✓ Worked example 2: the route on the strip

$$S_1/P_1 = S_2/P_2$$

given: 5.80 g/L at 4.0 atm · $P_2 = 1.0$ atm · found: $S_2 = 1.45$ g/L · 4.35 g released per liter



 **Sense check:** The top lane gives the new solubility, 1.45 g/L. The bottom lane turns the drop into gas released: $(5.80 - 1.45) \text{ g/L} \times 1 \text{ L} = 4.35 \text{ g}$. ✓

Your turn: nitrogen under pressure

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 0.019 \text{ g/L}$ · $P_1 = 1.0 \text{ atm}$ · $P_2 = 5.0 \text{ atm}$ · same gas, constant temperature · wanted: S_2

Nitrogen dissolves at **0.019 g/L** under **1.0 atm** of N_2 . A reactor blankets it at **5.0 atm**.

$$S_2 = 0.019 \text{ g/L} \times \frac{\text{ } \text{atm}}{\text{ } \text{atm}} = \text{ } \text{g/L}$$

Put the new pressure on top and the old pressure below, then compute.

$$S_2 = 0.019 \text{ g/L} \times \frac{5.0 \text{ atm}}{1.0 \text{ atm}} = \mathbf{0.095 \text{ g/L}}$$

 **Sense check:** Pressure rose fivefold, so five times as much N_2 dissolves: $0.019 \rightarrow 0.095 \text{ g/L}$. ✓

⚠ Where this goes wrong

$$S_1/P_1 = S_2/P_2$$

1.45 g/L of CO₂ at 1.0 atm, raised to 4.0 atm · correct $S_2 = 5.80$ g/L

Inverting the proportion. Pressure rose from 1.0 to 4.0 atm, so more CO₂ dissolves. Writing $1.45 \times (1.0/4.0) = 0.363$ g/L predicts less: that is Boyle's inverse ratio. Henry's law is direct: the new pressure goes on top, (P_2/P_1) .

Adding the pressure change. Solubility is a proportion, not a sum. Adding the change, $1.45 + (4.0 - 1.0) = 4.45$ g/L, tacks atm onto g/L. Multiply by the pressure ratio instead.

Assuming no change. The partial pressure quadrupled, so the dissolved amount cannot stay 1.45 g/L. Raising the pressure forces more gas in; the solubility must rise.

 Practice 2

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 0.0430 \text{ g/L}$ · $P_1 = 1.00 \text{ atm}$ · $P_2 = 3.00 \text{ atm}$ · same gas, constant temperature · wanted: S_2

Oxygen dissolves at **0.0430 g/L** under **1.00 atm** of O_2 . Under a **3.00 atm** atmosphere of O_2 at the same temperature, what is its solubility, in g/L?

- A. 0.0143
- B. 0.0860
- C. 0.129
- D. 0.0430

✓ Practice 2 answer: C

$$S_1/P_1 = S_2/P_2$$

$$\text{O}_2 \cdot S_1 = 0.0430 \text{ g/L at } P_1 = 1.00 \text{ atm} \cdot P_2 = 3.00 \text{ atm} \cdot \text{constant temperature}$$

$$S_2 = 0.0430 \text{ g/L} \times \frac{3.00 \text{ atm}}{1.00 \text{ atm}} = \mathbf{0.129 \text{ g/L} \cdot \text{answer C}}$$

A inverted the proportion like Boyle's law: $0.0430 \times (1.00/3.00) = 0.0143 \text{ g/L}$, less dissolved though the pressure rose. B scaled by the pressure change: $0.0430 \times (3.00 - 1.00)/1.00 = 0.0860 \text{ g/L}$; the ratio takes the new pressure, 3.00 atm, not the 2.00 atm increase. D assumed no change: at 3.00 atm the solubility cannot stay 0.0430 g/L.

🧠 **Sense check:** Pressure tripled, so three times as much O_2 dissolves: $0.0430 \rightarrow 0.129 \text{ g/L}$. ✓



Worked example 3: the pressure to reach a target

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 1.45 \text{ g/L}$ at $P_1 = 1.0 \text{ atm}$ · target $S_2 = 8.70 \text{ g/L}$ · same gas, constant temperature · wanted: P_2

A bottler wants **8.70 g/L** of CO_2 dissolved. At **1.0 atm** only **1.45 g/L** dissolves. What CO_2 partial pressure reaches the target?

A common first attempt: invert the ratio, the way Boyle's law does. Test the result.

✓ Worked example 3: solution

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 1.45 \text{ g/L}$ at $P_1 = 1.0 \text{ atm}$ · target $S_2 = 8.70 \text{ g/L}$ · wanted: P_2

A common first attempt

$$P_2 = 1.0 \text{ atm} \times \frac{1.45 \text{ g/L}}{8.70 \text{ g/L}} = 0.167 \text{ atm} \times$$

Step 1 · Identify given and wanted

Step 2 · Set up the proportion

Same gas, constant temperature, so S/P stays constant: $S_1/P_1 = S_2/P_2$.

 **Sense check:** The solubility must rise sixfold, so the pressure must rise too, not fall. ✓

✓ Worked example 3: the pressure needed

$$S_1/P_1 = S_2/P_2$$

$$S_1 = 1.45 \text{ g/L} \cdot P_1 = 1.0 \text{ atm} \cdot S_2 = 8.70 \text{ g/L} \cdot \text{wanted: } P_2$$

Step 3 • Rearrange for the unknown

$$S_1/P_1 = S_2/P_2 \rightarrow P_2 = P_1 \times \frac{S_2}{S_1}$$

Step 4 • Substitute and check

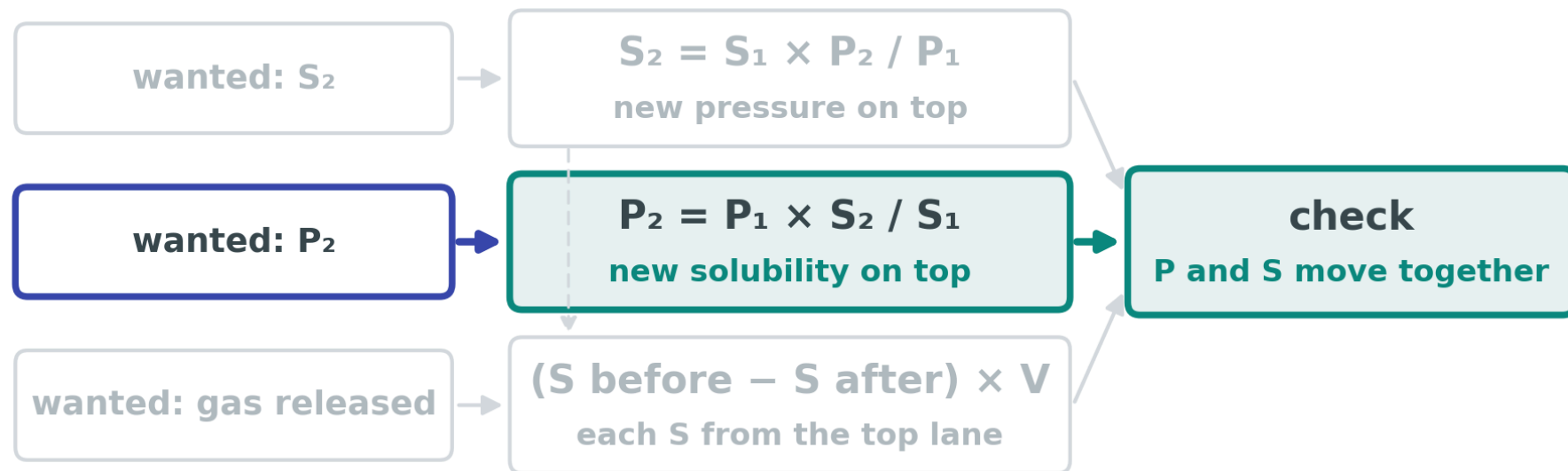
$$P_2 = 1.0 \text{ atm} \times \frac{8.70 \text{ g/L}}{1.45 \text{ g/L}} = \mathbf{6.0 \text{ atm}}$$

 **Sense check:** To dissolve six times as much CO₂, the partial pressure must be six times higher: 1.0 → 6.0 atm. ✓

✓ Worked example 3: the route on the strip

$$S_1/P_1 = S_2/P_2$$

given: 1.45 g/L at 1.0 atm · target $S_2 = 8.70$ g/L · found: $P_2 = 6.0$ atm



 **Sense check:** A pressure is wanted, so the middle lane: the new solubility goes on top. The solubility rose sixfold, and so did the pressure. ✓

Take-home: Henry's law is direct

inverted (Boyle's ratio): $1.0 \times (1.45 / 8.70) \rightarrow P = 0.167 \text{ atm}$

less pressure for more gas: impossible ✕

direct (Henry's law): $1.0 \times (8.70 / 1.45) \rightarrow P = 6.0 \text{ atm}$

more gas needs more pressure ✓

Solubility climbs with partial pressure, so the larger solubility goes on top of the ratio. Boyle's upside-down ratio would demand less pressure for more gas. That cannot happen.

 Practice 3

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 0.140 \text{ M}$ at $P_1 = 4.20 \text{ atm}$ · $S_2 = 0.0850 \text{ M}$ · same gas, constant temperature · wanted: P_2

A soda fountain's carbonator holds water at **0.140 M** CO_2 under **4.20 atm** of CO_2 . What CO_2 pressure, in atm, gives **0.0850 M** at the same temperature?

- A. 2.55
- B. 6.92
- C. 0.607
- D. 4.20

✓ Practice 3 answer: A

$$S_1/P_1 = S_2/P_2$$

$$\text{CO}_2 \cdot S_1 = 0.140 \text{ M at } P_1 = 4.20 \text{ atm} \cdot \text{target } S_2 = 0.0850 \text{ M} \cdot \text{wanted: } P_2$$

$$P_2 = 4.20 \text{ atm} \times \frac{0.0850 \text{ M}}{0.140 \text{ M}} = 2.55 \text{ atm} \cdot \text{answer A}$$

B inverted the ratio like Boyle's law: $4.20 \times (0.140/0.0850) = 6.92 \text{ atm}$, more pressure for less dissolved gas. C stopped at the ratio: $0.0850/0.140 = 0.607$, never multiplied by P_1 . D assumed no change: a lower concentration needs a lower pressure.

 **Sense check:** The concentration falls to 0.607 of its value, so the pressure falls to $0.607 \times 4.20 \text{ atm} = 2.55 \text{ atm}$. ✓



Practice 4

$$S_1/P_1 = S_2/P_2$$

given: CH₄ 0.0460 g/L at 2.00 atm · saturated at 6.00 atm · released to 1.50 atm · 400. mL · wanted: g CH₄ released

Deep well water is saturated with methane under **6.00 atm** of CH₄. At the wellhead the CH₄ pressure falls to **1.50 atm** at the same temperature. Methane dissolves at 0.0460 g/L under 2.00 atm. How many grams of CH₄ leave a **400. mL** sample?

- A. 0.0552
- B. 0.0414
- C. 0.0828
- D. 41.4

✓ Practice 4 answer: B

$$S_1/P_1 = S_2/P_2$$

$$\text{CH}_4 \text{ 0.0460 g/L at 2.00 atm} \cdot 6.00 \text{ atm} \rightarrow 1.50 \text{ atm} \cdot 400. \text{ mL} = 0.400 \text{ L}$$

$$S_{6.00} = 0.0460 \text{ g/L} \times \frac{6.00}{2.00} = 0.138 \text{ g/L} \cdot S_{1.50} = 0.0460 \text{ g/L} \times \frac{1.50}{2.00} = 0.0345 \text{ g/L}$$

$$(0.138 - 0.0345) \frac{\text{g}}{\text{L}} \times 0.400 \text{ L} = \mathbf{0.0414 \text{ g CH}_4 \cdot \text{answer B}}$$

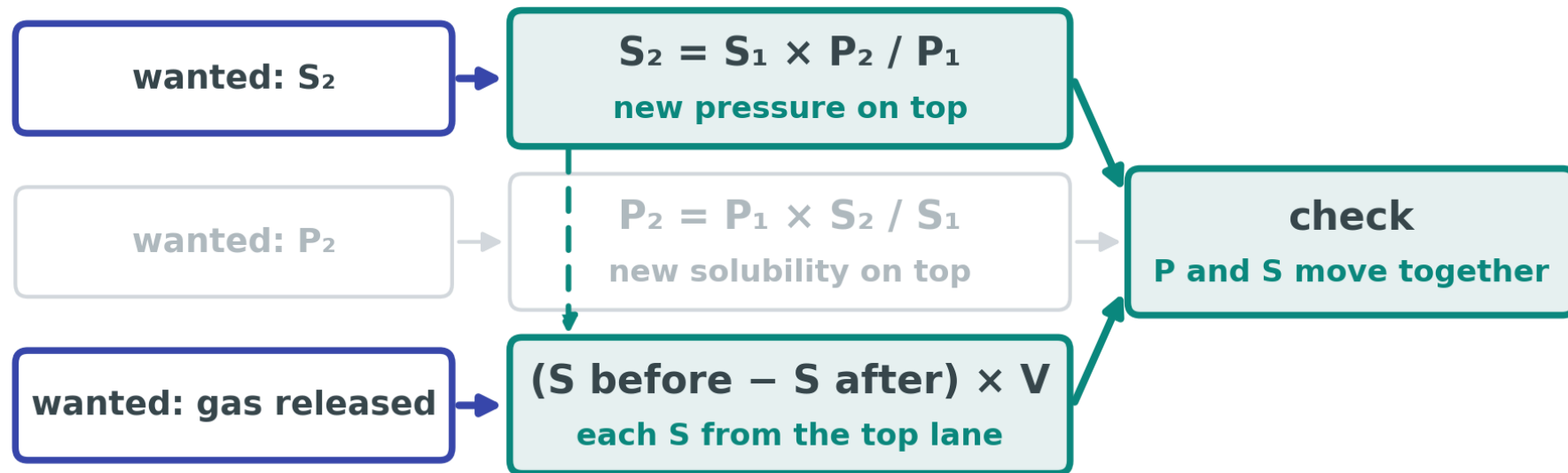
A is all the CH₄ dissolved at 6.00 atm: 0.138 × 0.400 = 0.0552 g. C skipped the ÷ 2.00 atm: 0.0828 g. D left mL unconverted: 41.4 g.

 **Sense check:** Pressure fell to a quarter, so three quarters leaves ✓

✓ Practice 4: the route on the strip

$$S_1/P_1 = S_2/P_2$$

given: CH₄ 0.0460 g/L at 2.00 atm · 6.00 atm → 1.50 atm · 0.400 L · found: 0.0414 g released



 **Sense check:** Two trips through the top lane give each solubility: 0.138 g/L at 6.00 atm and 0.0345 g/L at 1.50 atm. The bottom lane turns the 0.1035 g/L drop into grams: $\times 0.400 \text{ L} = 0.0414 \text{ g}$. ✓



Practice 5

$$S_1/P_1 = S_2/P_2$$

given: $S_1 = 8.40 \text{ mg/L}$ at $P_1 = 159 \text{ mmHg}$ · target $S_2 = 12.0 \text{ mg/L}$ · same gas, constant temperature ·
wanted: P_2

Under air, a trout hatchery's water holds **8.40 mg/L** of O_2 at an O_2 pressure of **159 mmHg**. The fish need **12.0 mg/L**. What O_2 partial pressure, in mmHg, reaches that target at the same temperature?

- A. 111
- B. 1.43
- C. 163
- D. 227

✓ Practice 5 answer: D

$$S_1/P_1 = S_2/P_2$$

$$O_2 \cdot S_1 = 8.40 \text{ mg/L at } P_1 = 159 \text{ mmHg} \cdot \text{target } S_2 = 12.0 \text{ mg/L} \cdot \text{wanted: } P_2$$

$$P_2 = 159 \text{ mmHg} \times \frac{12.0 \text{ mg/L}}{8.40 \text{ mg/L}} = \mathbf{227 \text{ mmHg} \cdot \text{answer D}}$$

A inverted the ratio like Boyle's law: $159 \times (8.40/12.0) = 111 \text{ mmHg}$, less pressure for more dissolved O_2 . B stopped at the ratio: $12.0/8.40 = 1.43$, never multiplied by P_1 . C added the solubility change to the pressure: $159 + (12.0 - 8.40) = 163$, which mixes mg/L with mmHg.

 **Sense check:** The target is 1.43 times the O_2 the water holds now, so the pressure rises by the same factor: $1.43 \times 159 \text{ mmHg} = 227 \text{ mmHg}$. ✓

 Practice 6

$$S_1/P_1 = S_2/P_2$$

given: CO₂ 2.77 g/L at 1.00 atm · saturated at 3.40 atm · vented to 2.00 atm · 946 mL · wanted: g CO₂ released

A seltzer keg at 5 °C is saturated under **3.40 atm** of CO₂, then vented until the CO₂ pressure is **2.00 atm**. At 5 °C, CO₂ dissolves to 2.77 g/L under 1.00 atm. How many grams of CO₂ leave **946 mL** of the seltzer?

- A. 3.67
- B. 8.91
- C. 3.88
- D. 5.24

✓ Practice 6 answer: A

$$S_1/P_1 = S_2/P_2$$

$$\text{CO}_2 \text{ 2.77 g/L at 1.00 atm} \cdot 3.40 \text{ atm} \rightarrow 2.00 \text{ atm} \cdot 946 \text{ mL} = 0.946 \text{ L}$$

$$S_{3.40} = 2.77 \text{ g/L} \times \frac{3.40}{1.00} = 9.418 \text{ g/L} \cdot S_{2.00} = 2.77 \text{ g/L} \times \frac{2.00}{1.00} = 5.540 \text{ g/L}$$

$$(9.418 - 5.540) \frac{\text{g}}{\text{L}} \times 0.946 \text{ L} = 3.67 \text{ g CO}_2 \cdot \text{answer A}$$

B is all the CO₂ dissolved at 3.40 atm: $9.418 \times 0.946 = 8.91 \text{ g}$. C stopped at the drop per liter, 3.88 g/L, and never multiplied by 0.946 L. D is the CO₂ that stays dissolved at 2.00 atm: $5.540 \times 0.946 = 5.24 \text{ g}$.

 **Sense check:** The pressure fell by 1.40 atm, so each liter loses $1.40 \times 2.77 = 3.878 \text{ g}$, and 0.946 L loses 3.67 g. ✓

 Practice 7

$$S_1/P_1 = S_2/P_2 \cdot \text{mol} = M \times L$$

given: CO₂ 0.0336 M at 1.00 atm · bottled under 3.50 atm · 591 mL · wanted: mol CO₂ dissolved

At 25 °C, CO₂ dissolves to **0.0336 M** under 1.00 atm of CO₂. A soda is bottled under **3.50 atm** of CO₂ at the same temperature. How many moles of CO₂ are dissolved in a **591 mL** bottle?

- A. 0.0199
- B. 0.118
- C. 0.0695
- D. 0.00567

✓ Practice 7 answer: C

$$S_1/P_1 = S_2/P_2 \cdot \text{mol} = M \times L$$

$$\text{CO}_2 \text{ 0.0336 M at 1.00 atm} \cdot \text{bottled under 3.50 atm} \cdot 591 \text{ mL} = 0.591 \text{ L}$$

$$S_2 = 0.0336 \text{ M} \times \frac{3.50 \text{ atm}}{1.00 \text{ atm}} = 0.1176 \text{ M}$$

$$0.591 \text{ L} \times \frac{0.1176 \text{ mol}}{1 \text{ L}} = 0.0695 \text{ mol CO}_2 \cdot \text{answer C}$$

🧠 **Sense check:** 3.50 times the pressure dissolves 3.50 times the CO₂: 0.01986 mol × 3.50 = 0.0695 mol. ✓

Check yourself

1. A sealed soda is opened and the gas pressing above it drops sharply. Does the dissolved gas rise or fall, and why does the drink fizz?
2. Write $S_1/P_1 = S_2/P_2$ solved for S_2 . When the partial pressure triples, what happens to the solubility?

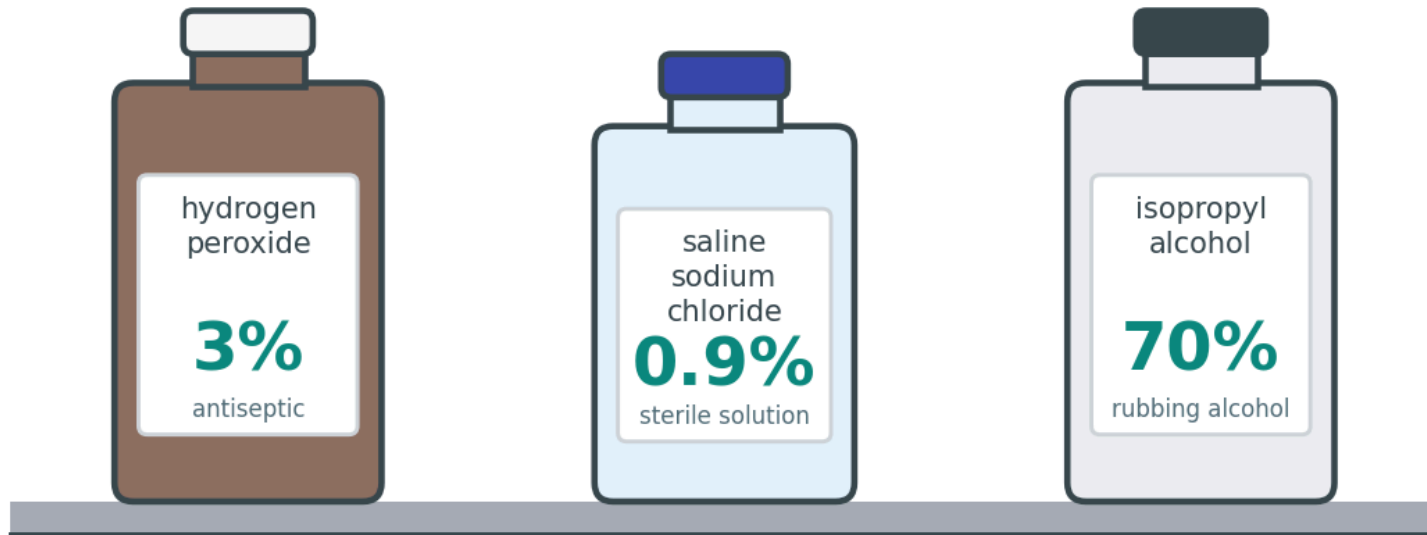
Henry's law gives a gas's concentration in g/L, mg/L or M. Percent concentration describes any solution's makeup a different way: grams of solute per 100 g of solution.

6 • 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.

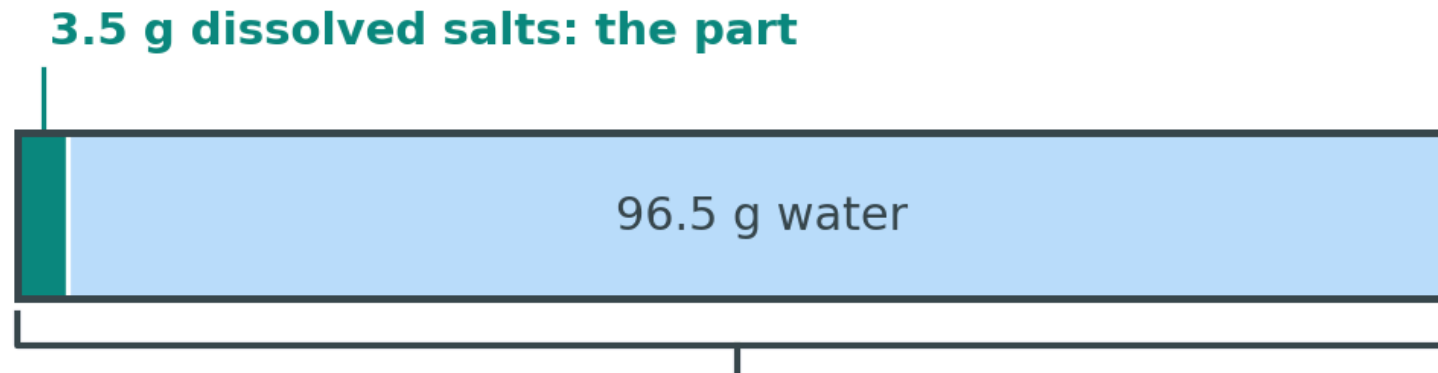
Which drink is more concentrated: 10.0 g of sugar in 200. g of drink, or 15.0 g in 500. g?

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

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:

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

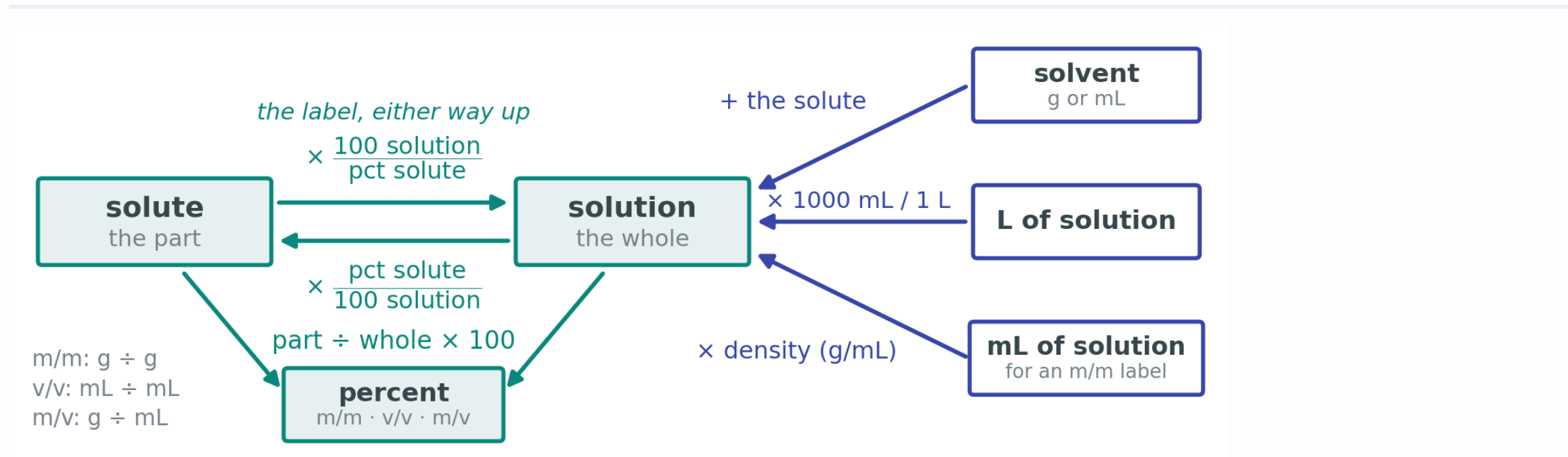
Write it so the **given unit cancels**.



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

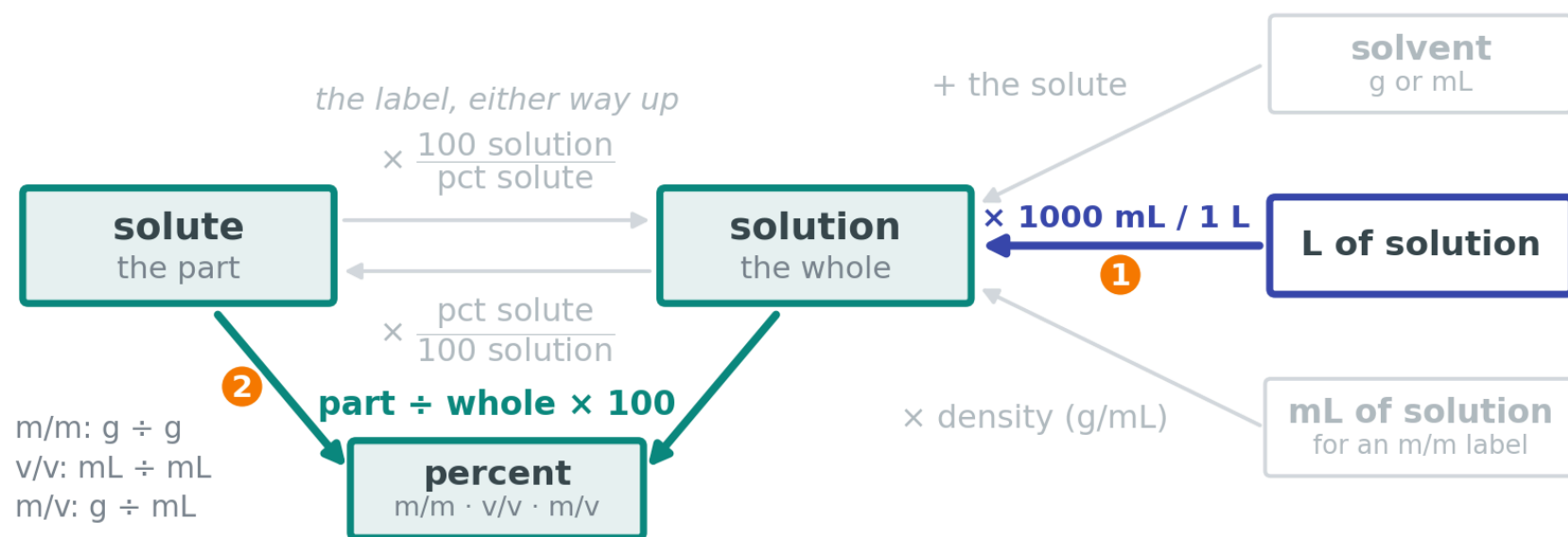
Step 4 · Multiply and check

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

 **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

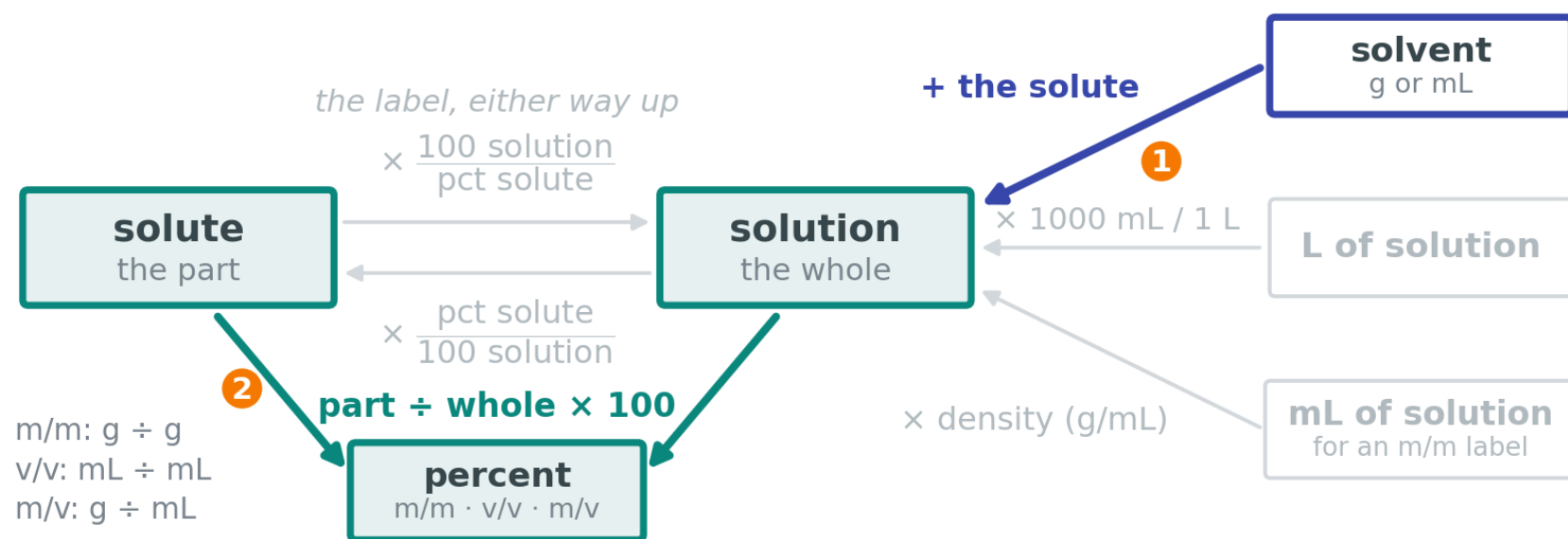
$$\frac{25.0 \text{ g sucrose}}{125.0 \text{ g solution}} \times 100 = \mathbf{20.0\% (m/m)}$$

 **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

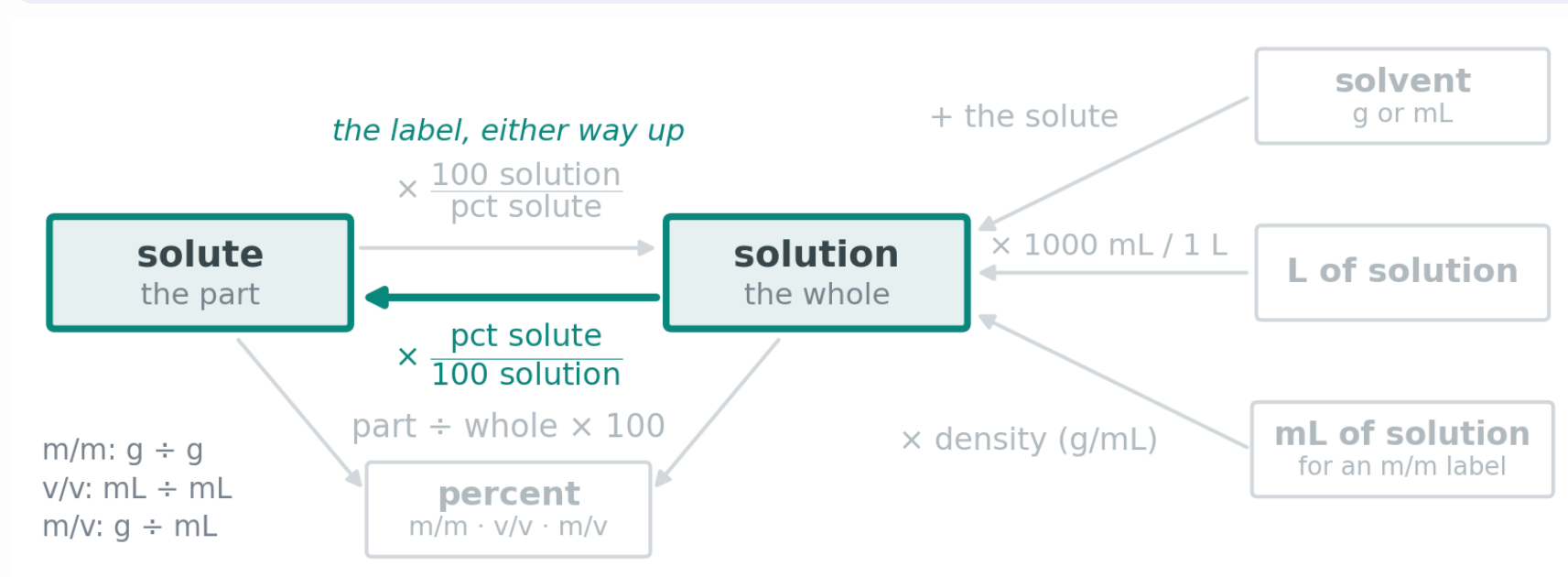
Step 4 · Multiply and check

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

 **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

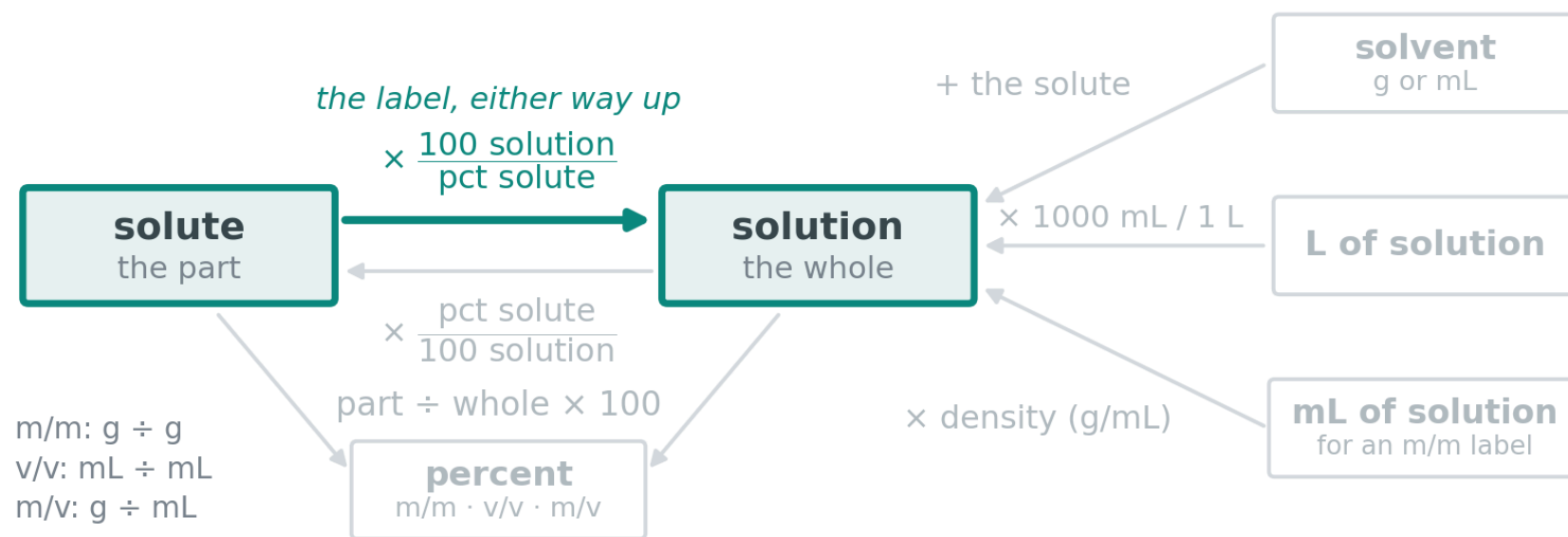
Step 4 · Multiply and check

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

 **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.

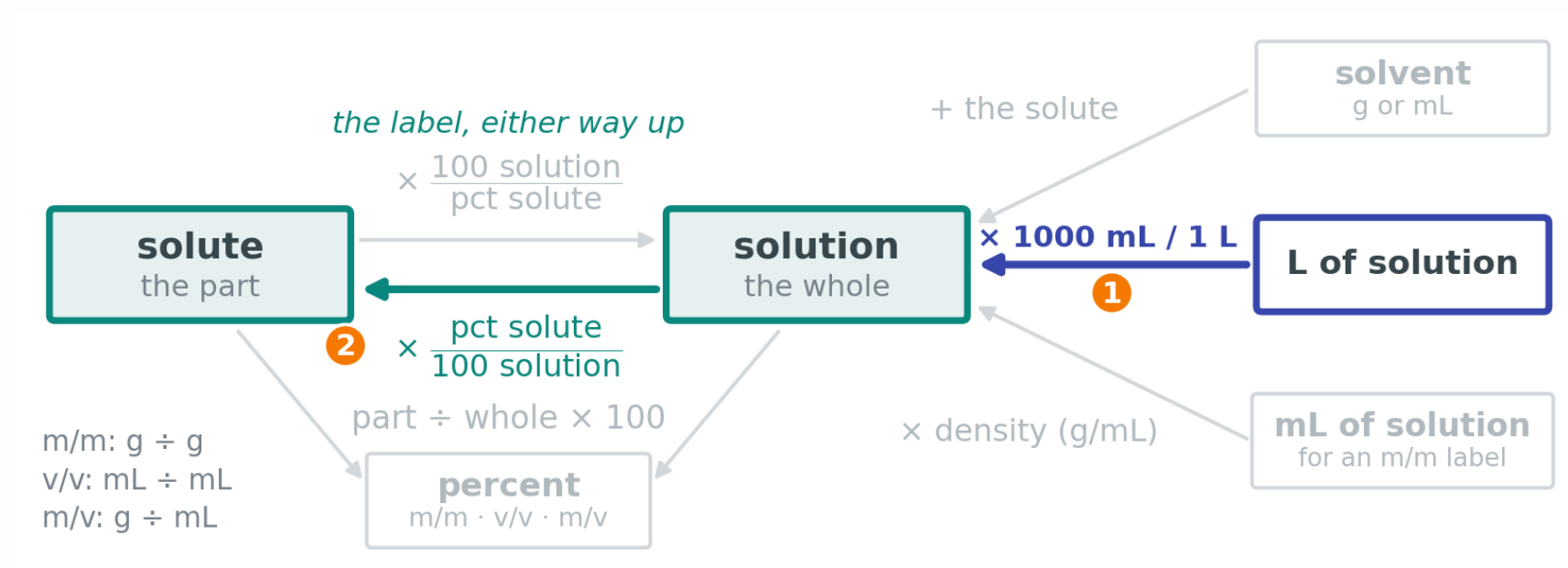
$$0.500 \text{ L soln} \times \frac{1000 \text{ mL soln}}{1 \text{ L soln}} \times \frac{15.0 \text{ g mannitol}}{100 \text{ mL soln}} = \mathbf{75.0 \text{ g mannitol (answer D)}}$$

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 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

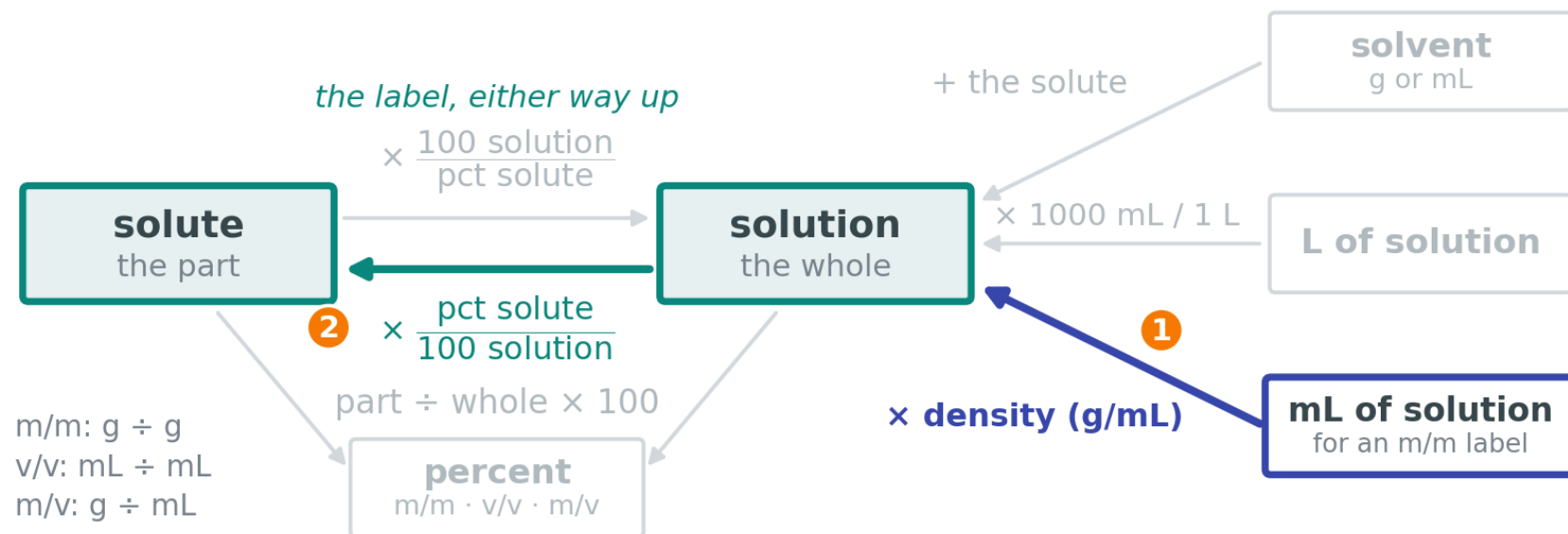
$$250. \cancel{\text{ mL soln}} \times \frac{0.900 \cancel{\text{ g soln}}}{1 \cancel{\text{ mL soln}}} \times \frac{28.0 \text{ g NH}_3}{100 \cancel{\text{ g soln}}} = 63.0 \text{ g NH}_3 \text{ (answer B)}$$

A skipped the density: 250. × 28.0 / 100 = 70.0 treats mL of solution as grams. C stopped halfway: 250. × 0.900 = 225 g is the whole solution, not the NH₃ in it. D flipped the density: 250. ÷ 0.900 = 278 g, then × 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. ✓

✓ 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.

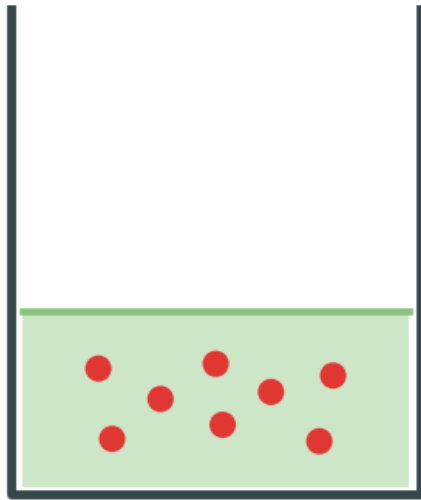
7 · Molality & Mole Fraction

Calculate the molality of a solution from grams of solute and grams of solvent and explain why it does not change with temperature.

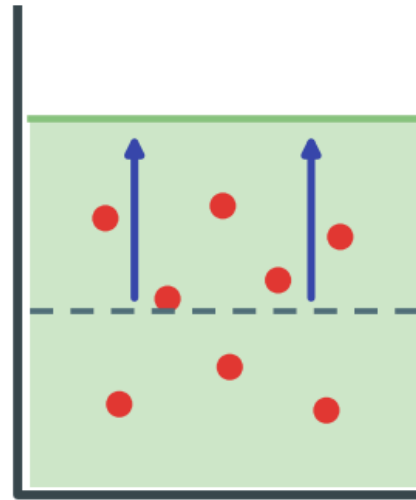
Coolant that warms and swells

Warm coolant expands and rises

the antifreeze in it — the same eight particles, the same weight — never moved



a cold engine
coolant sits low



running hot
the same coolant, risen

Antifreeze in a cold engine sits at a low line. Running hot, the same liquid expands and rises. Its blend never changed: only the space it fills.

Three concentration units, three denominators

mass percent = g solute ÷ g solution × 100

molarity, M = mol solute ÷ L solution

molality, m = mol solute ÷ kg solvent

solution = solute + solvent · the solute always sits on top

Every concentration puts the solute on top. The denominators differ: the solution's mass, the solution's volume, or the solvent's mass.

Which denominator leaves the solute out?

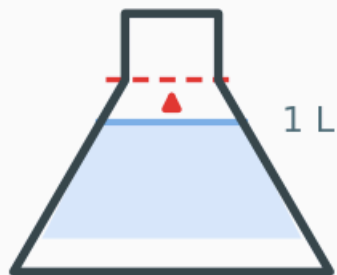
Molality's: the solvent alone, in kilograms.

Measured by amount and mass, not volume

Size a solution by its volume, or by its mass

molarity

moles of solute $\div$ liters of SOLUTION

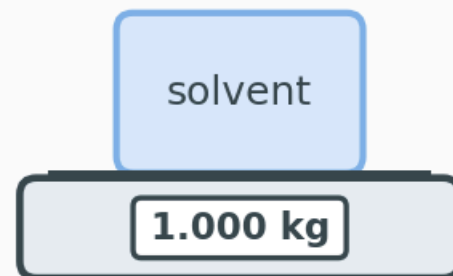


warm it — the volume rises

VS

molality

moles of solute $\div$ kilograms of SOLVENT



warm it — the mass holds

Molarity divides by the solution's volume, and volume swells when warmed. Molality counts moles and weighs mass, so temperature does not change it.

Molality: moles of solute per kilogram of solvent

$$m = \text{mol solute} \div \text{kg solvent}$$

a 2.0 m solution: every kilogram of the SOLVENT carries 2.0 mol solute · read "two molal"

Molality, symbol m , counts the moles of solute in each kilogram of solvent. The denominator is the solvent alone: not the solution, and not a volume.

memory hook: little m , mass of solvent · big M , liters of solution

$$m = \text{mol} \div \text{kg solvent} \cdot M = \text{mol} \div \text{L solution}$$

A molality is an equality

1.50 m glucose: 1.50 mol glucose = 1 kg water
every kilogram of water carries 1.50 mol of glucose

Every equality gives a conversion factor. This one links moles of solute to kilograms of solvent:

$$\frac{1.50 \text{ mol glucose}}{1 \text{ kg water}} \text{ or } \frac{1 \text{ kg water}}{1.50 \text{ mol glucose}}$$

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

Temperature changes volume, not mass

warm a solution: volume 1.00 L $\rightarrow$ 1.03 L · molarity falls
the moles are unchanged, but they now spread through more liters

warm the same solution: solvent mass 1.00 kg $\rightarrow$ 1.00 kg · molality holds
mass does not expand, so moles per kilogram stay put

Heat expands a liquid's volume but never its mass. Molarity changes with temperature; molality does not.

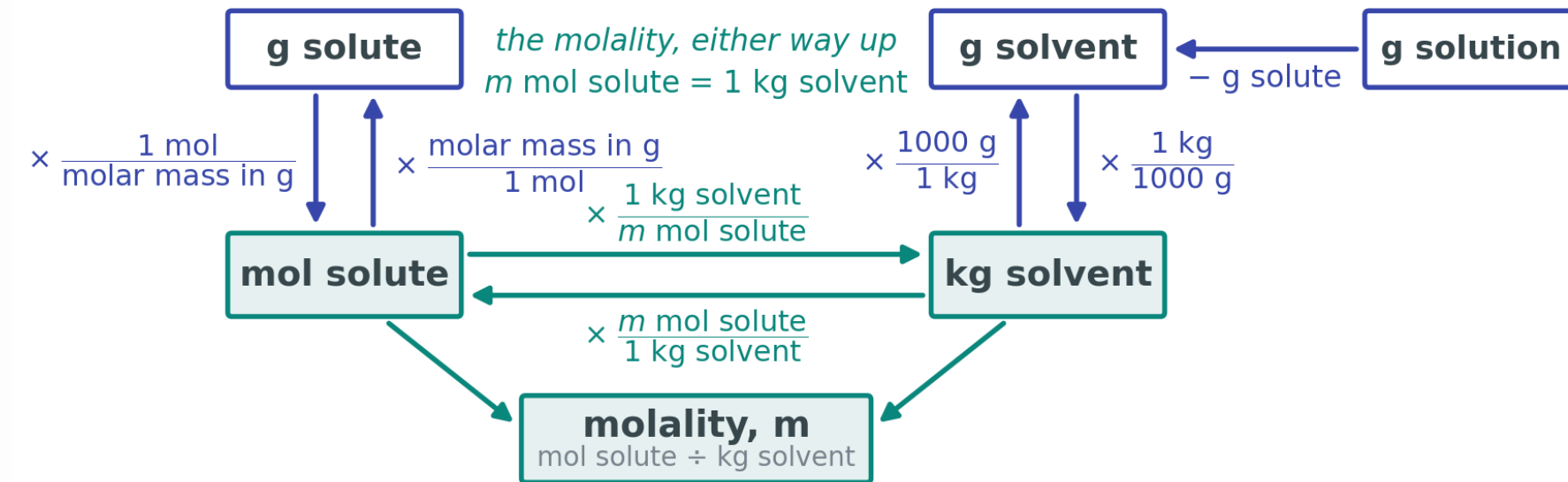


The method

1. **Grams $\rightarrow$ moles:** convert the solute's mass with its molar mass.
2. **Grams $\rightarrow$ kilograms:** divide the solvent's mass by 1000.
3. **Divide moles by kilograms:** the quotient is the molality, in mol/kg.

Backward: kilograms $\times$ molality $\rightarrow$ moles $\rightarrow$ grams.

One map for every molality problem



Forward: solute grams become moles, solvent grams become kilograms, then moles $\div$ kilograms. Backward: the molality, either way up, crosses between the two. A total solution mass loses its solute first.



Guided example: brine for an ice-cream maker

Step 1 • Grams → moles

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 1.20 mol NaCl · 800. g water · wanted: m

An old-fashioned ice-cream maker packs its tub in a brine of **1.20 mol of NaCl** dissolved in **800. g of water**. Find the molality.

The salt is already counted in moles, so no molar mass is needed.

✓ Guided example: solution

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 1.20 mol NaCl · 800. g water · wanted: m

Step 2 · Grams → kilograms

$$800. \text{ g water} = 0.800 \text{ kg water}$$

1000 g = 1 kg · molality counts kilograms of the solvent

Step 3 · Divide moles by kilograms

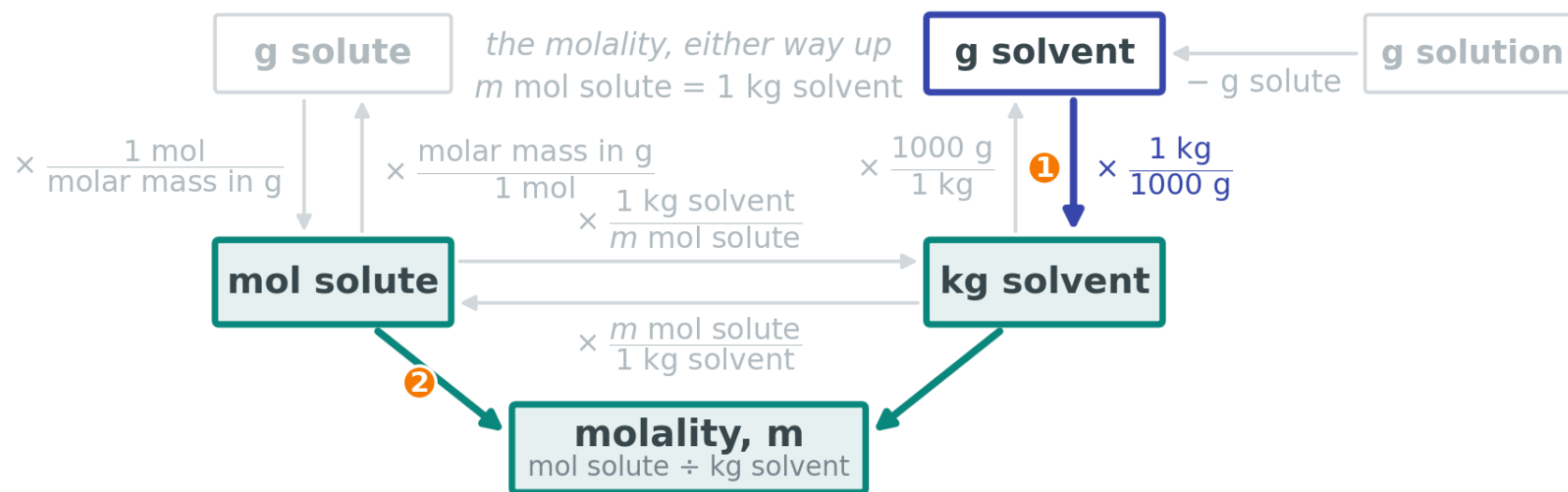
$$m = \frac{1.20 \text{ mol NaCl}}{0.800 \text{ kg water}} = 1.50 \text{ mol/kg}$$

🧠 **Sense check:** 0.800 kg of water holds 1.20 mol, so a full kilogram holds more: 1.50 mol. 1.50 m ✓

✓ Guided example: the route on the map

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 1.20 mol NaCl · 800. g water · found: 1.50 mol/kg



🧠 **Sense check:** The salt came in moles, so no molar mass. Two moves: grams to kilograms, then moles ÷ kilograms.





Practice 1

fruit syrup: fructose ($\text{C}_6\text{H}_{12}\text{O}_6$) dissolved in water

given: 0.840 mol fructose · 600. g water · wanted: m

A fruit syrup holds **0.840 mol of fructose** in **600. g of water**. What is its molality?

- A. 0.504
- B. 1.40×10^{-3}
- C. 1.40
- D. 0.714

✓ Practice 1 answer: C

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 0.840 mol fructose · 600. g water · wanted: m

$$600. \text{ g water} = 0.600 \text{ kg water} \cdot m = \frac{0.840 \text{ mol fructose}}{0.600 \text{ kg water}} = \mathbf{1.40 \text{ mol/kg} \cdot \text{answer C}}$$

A multiplied instead of dividing: $0.840 \times 0.600 = 0.504$. B divided by grams of water: $0.840 \div 600 = 1.40 \times 10^{-3}$, 1000 times too small. D flipped the fraction: $0.600 \div 0.840 = 0.714$ counts kilograms per mole.

 **Sense check:** More moles than kilograms, so the molality tops 1: 0.600 kg carries 0.840 mol, a full kilogram 1.40 mol ✓



Worked example 1: molality from grams

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 18.0 g glucose ($\text{C}_6\text{H}_{12}\text{O}_6$, 180.16 g/mol) · 250. g water · wanted: m

A syrup is made by dissolving **18.0 g of glucose** ($\text{C}_6\text{H}_{12}\text{O}_6$, 180.16 g/mol) in **250. g of water**. Find the molality.

Set it up: turn the solute into moles and the solvent into kilograms, then divide.

✓ Worked example 1: solution

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 18.0 g glucose (180.16 g/mol) · 250. g water · wanted: m

Step 1 · Grams → moles

Step 2 · Grams → kilograms

$$250. \text{ g water} = 0.250 \text{ kg water}$$

1000 g = 1 kg · molality counts kilograms of the solvent

Step 3 · Divide moles by kilograms

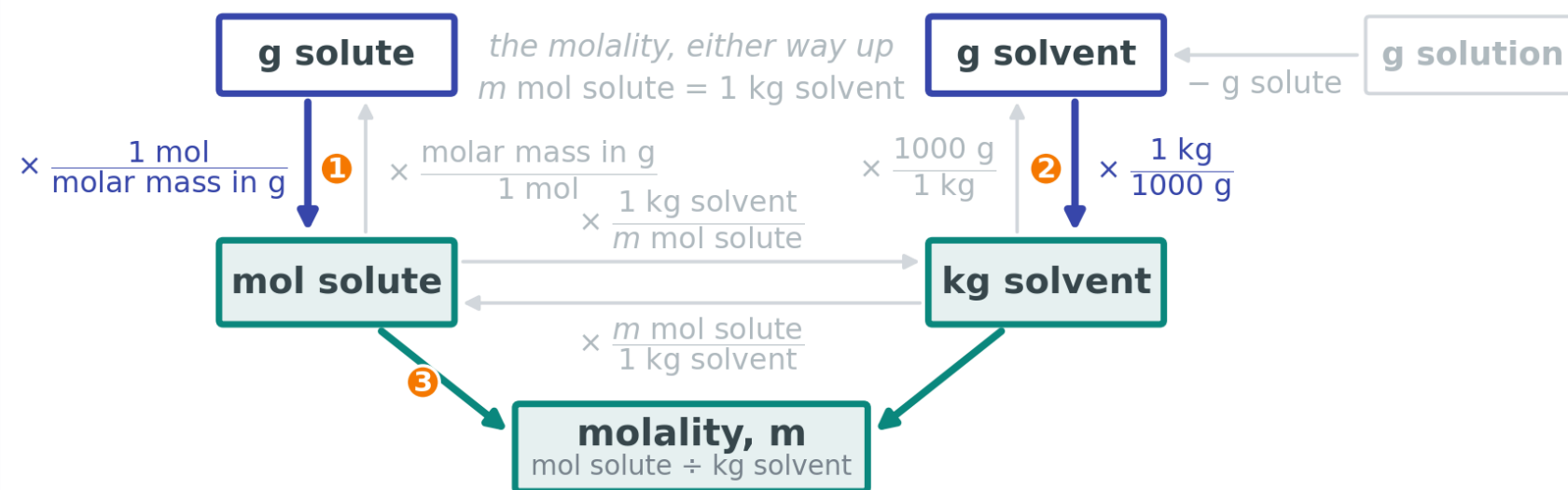
$$m = \frac{0.0999 \text{ mol glucose}}{0.250 \text{ kg water}} = 0.400 \text{ mol/kg}$$

 **Sense check:** A quarter kilogram of water carries 0.0999 mol, so a full kilogram carries four times that: 0.400 mol.
0.400 m ✓

✓ Worked example 1: the route on the map

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 18.0 g glucose (180.16 g/mol) · 250. g water · found: 0.400 mol/kg



🧠 **Sense check:** Three moves: each mass converts on its own side, then moles ÷ kilograms. ✓

Your turn: sucrose in water

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 34.2 g sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$, 342.30 g/mol) · 500. g water · wanted: m

A sweetened solution holds **34.2 g of sucrose** (342.30 g/mol) in **500. g of water**.

$$34.2 \text{ g sucrose} \times \frac{1 \text{ mol sucrose}}{\text{g sucrose}} = \text{mol}, \text{ then } \frac{\text{mol sucrose}}{\text{kg water}} = \text{mol/kg}$$

Fill the molar mass, the moles, and the kilograms of solvent, then compute the molality.

$$34.2 \text{ g sucrose} \times \frac{1 \text{ mol sucrose}}{342.30 \text{ g sucrose}} = 0.0999 \text{ mol}, \text{ then } \frac{0.0999 \text{ mol sucrose}}{0.500 \text{ kg water}} = 0.200 \text{ mol/kg}$$

⚠ Where this goes wrong

18.0 g glucose (180.16 g/mol) · 250. g water
correct: 18.0 g → 0.0999 mol · 250. g → 0.250 kg · m = 0.400 mol/kg

Dividing by grams of solvent. $0.0999 \text{ mol} \div 250 \text{ g} = 0.000400$, and the unit is mol/g, 1000 times too small. Molality divides by kilograms: $0.0999 \div 0.250 = 0.400 \text{ mol/kg}$.

Using the mass of the solution. The whole solution is $250 + 18.0 = 268 \text{ g}$, or 0.268 kg , giving $0.0999 \div 0.268 = 0.373$. Molality uses the solvent alone: 0.250 kg .

Skipping the molar mass. $18.0 \text{ g} \div 0.250 \text{ kg} = 72.0$ treats grams as moles. Grams become moles first: $18.0 \div 180.16 = 0.0999 \text{ mol}$.



Practice 2

hand-lotion base: glycerol ($\text{C}_3\text{H}_8\text{O}_3$) dissolved in water

molar mass: glycerol 92.09 g/mol

A lotion base holds **46.0 g of glycerol** in **400. g of water**. What is the molality of the glycerol?

- A. 1.25
- B. 1.12
- C. 115
- D. 0.00125

✓ Practice 2 answer: A

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 46.0 g glycerol (92.09 g/mol) · 400. g water · wanted: m

$$46.0 \text{ g glycerol} \times \frac{1 \text{ mol glycerol}}{92.09 \text{ g glycerol}} = 0.500 \text{ mol} \cdot 400. \text{ g water} = 0.400 \text{ kg water}$$

$$m = \frac{0.500 \text{ mol glycerol}}{0.400 \text{ kg water}} = 1.25 \text{ mol/kg} \cdot \text{answer A}$$

B used the solution's mass: $0.500 \text{ mol} \div 0.446 \text{ kg} = 1.12$. C skipped the molar mass: $46.0 \text{ g} \div 0.400 \text{ kg} = 115$. D divided by grams of water: $0.500 \div 400 = 0.00125 \text{ mol/g}$, 1000 times too small.

 **Sense check:** 0.400 kg of water carries 0.500 mol, so a full kilogram carries 2.5 times that: 1.25 mol ✓



Worked example 2: grams of solute for a target molality

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 3.00 m urea ($\text{CH}_4\text{N}_2\text{O}$, 60.06 g/mol) · 250. g water · wanted: g urea

A fertilizer solution must be **3.00 m urea**, made with **250. g of water**. What mass of urea dissolves?

A common first attempt keeps the water in grams. Test the size of the answer.

✓ Worked example 2: solution

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 3.00 m urea (60.06 g/mol) · 250. g water · wanted: g urea

A common first attempt

$$250 \text{ g water} \times \frac{3.00 \text{ mol urea}}{1 \text{ kg water}} \times \frac{60.06 \text{ g urea}}{1 \text{ mol urea}} = 4.50 \times 10^4 \text{ g} \times$$

Backward · kilograms

Backward · molality → moles → grams

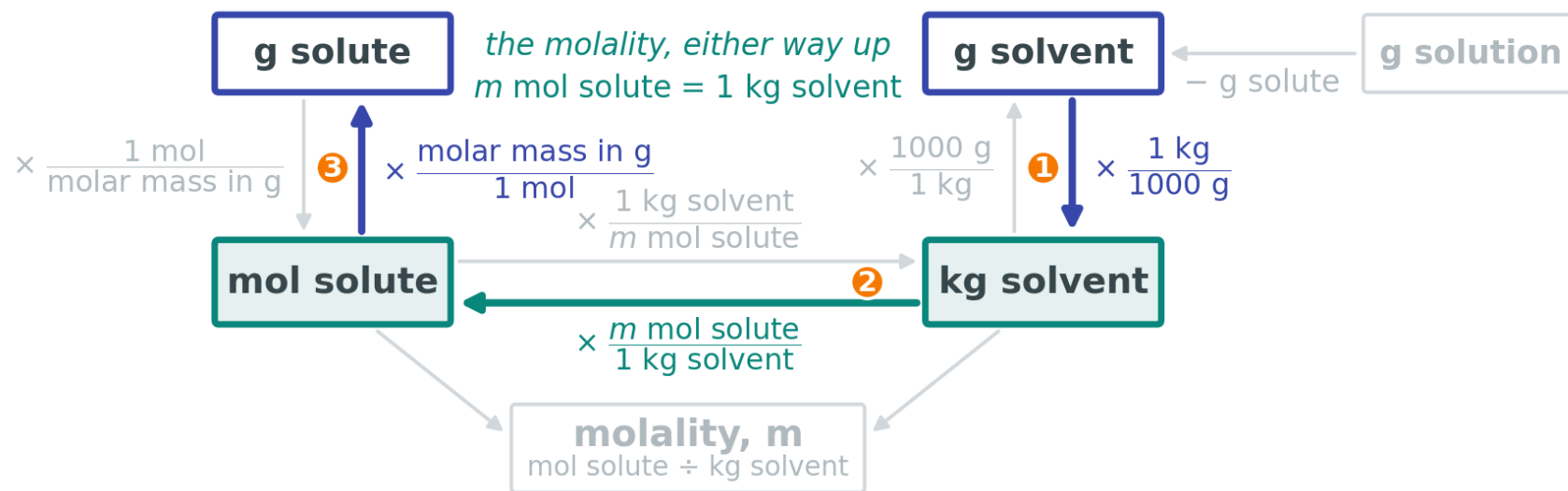
$$0.250 \text{ kg water} \times \frac{3.00 \text{ mol urea}}{1 \text{ kg water}} \times \frac{60.06 \text{ g urea}}{1 \text{ mol urea}} = 45.0 \text{ g urea}$$

 **Sense check:** A quarter kilogram of water takes a quarter of 3.00 mol: 0.750 mol, and 0.750 mol × 60.06 g/mol = 45.0 g ✓

✓ Worked example 2: the route on the map

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 3.00 m urea (60.06 g/mol) · 250. g water · found: 45.0 g urea



🧠 **Sense check:** Backward: the water enters in kilograms, the molality carries it across to moles, and the molar mass turns moles into grams. ✓

Take-home: the solvent enters in kilograms

$$250 \text{ g water} \times \frac{3.00 \text{ mol}}{1 \text{ kg water}} \times \frac{60.06 \text{ g}}{1 \text{ mol}} = 4.50 \times 10^4 \text{ g} \times \cdot \text{g and kg never cancel}$$

$$0.250 \text{ ~~kg water~~} \times \frac{3.00 \text{ mol}}{1 \text{ ~~kg water~~}} \times \frac{60.06 \text{ g}}{1 \text{ mol}} = \mathbf{45.0 \text{ g} \checkmark}$$

Molality is defined per kilogram of solvent. Solvent left in grams puts the answer off by 1000. Convert grams to kilograms, and use the solvent, never the solution.



Practice 3

de-icing fluid: propylene glycol ($\text{C}_3\text{H}_8\text{O}_2$) in water

molar mass: propylene glycol 76.09 g/mol

A de-icing fluid must be **1.60 m propylene glycol**, made with **750. g of water**. What mass of propylene glycol dissolves, in grams?

- A. 1.20
- B. 91.3
- C. 35.7
- D. 9.13×10^4

✓ Practice 3 answer: B

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 1.60 m propylene glycol (76.09 g/mol) · 750. g water · wanted: g propylene glycol

$$0.750 \text{ kg water} \times \frac{1.60 \text{ mol}}{1 \text{ kg water}} \times \frac{76.09 \text{ g}}{1 \text{ mol}} = 91.3 \text{ g} \cdot \text{answer B}$$

A stopped at moles: $0.750 \times 1.60 = 1.20 \text{ mol}$. C divided the kilograms by the molality: $0.750 \div 1.60 = 0.469 \text{ mol}$, then $\times 76.09 = 35.7$. D kept the water in grams: $750 \times 1.60 \times 76.09 = 9.13 \times 10^4$.

 **Sense check:** Three quarters of a kilogram takes three quarters of 1.60 mol: 1.20 mol, and $1.20 \text{ mol} \times 76.09 \text{ g/mol} = 91.3 \text{ g}$ ✓



Practice 4

aqueous sodium hydroxide (NaOH)

molar mass: NaOH 40.00 g/mol

How many grams of water dissolve **8.00 g of NaOH** to make a **0.500 m** solution?

- A. 100.
- B. 1.60×10^4
- C. 0.400
- D. 400.

✓ Practice 4 answer: D

0.500 m NaOH: 0.500 mol NaOH = 1 kg water

given: 8.00 g NaOH (40.00 g/mol) · wanted: g water · three conversion factors

$$8.00 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} = 0.200 \text{ mol NaOH}$$

$$0.200 \text{ mol NaOH} \times \frac{1 \text{ kg water}}{0.500 \text{ mol NaOH}} \times \frac{1000 \text{ g water}}{1 \text{ kg water}} = 400. \text{ g water} \cdot \text{answer D}$$

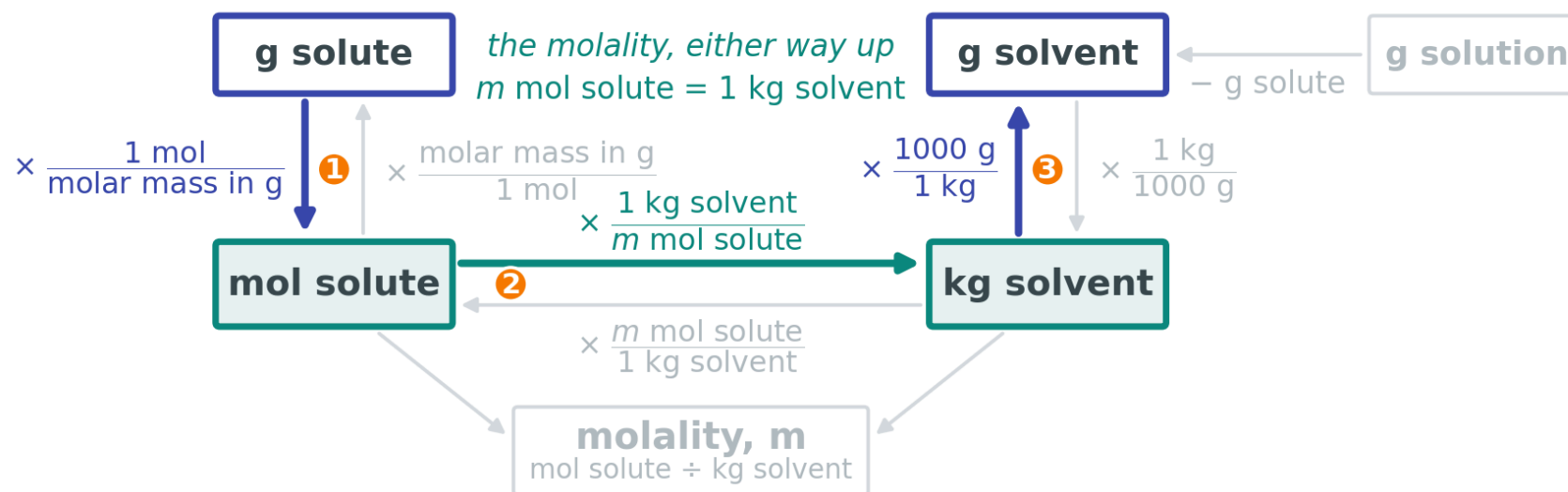
A flipped the molality factor: $0.200 \times 0.500 = 0.100 \text{ kg}$, or 100. g. B skipped the molar mass: $8.00 \div 0.500 = 16.0 \text{ kg}$, or $1.60 \times 10^4 \text{ g}$. C stopped at 0.400 kg.

 **Sense check:** A full kilogram carries 0.500 mol, so 0.200 mol needs 0.400 kg: 400. g ✓

✓ Practice 4: the route on the map

0.500 m NaOH: 0.500 mol NaOH = 1 kg water

given: 8.00 g NaOH · found: 400. g water



🧠 **Sense check:** Three moves: grams to moles, the molality flipped to carry moles to kilograms, then kilograms to grams. ✓



Practice 5

aqueous potassium chloride (KCl)

molar mass: KCl 74.55 g/mol

A solution holds **37.3 g of KCl** in enough water to make a total mass of **287.3 g**. What is the molality of the KCl?

- A. 1.74
- B. 0.500
- C. 1.54
- D. 2.00×10^{-3}
- E. 2.00

✓ Practice 5 answer: E

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 37.3 g KCl (74.55 g/mol) · 287.3 g solution · wanted: m

$$287.3 \text{ g solution} - 37.3 \text{ g KCl} = 250.0 \text{ g water} = 0.2500 \text{ kg water}$$

$$37.3 \text{ g KCl} \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} = 0.500 \text{ mol} \cdot m = \frac{0.500 \text{ mol}}{0.2500 \text{ kg}} = \mathbf{2.00 \text{ mol/kg} \cdot \text{answer E}}$$

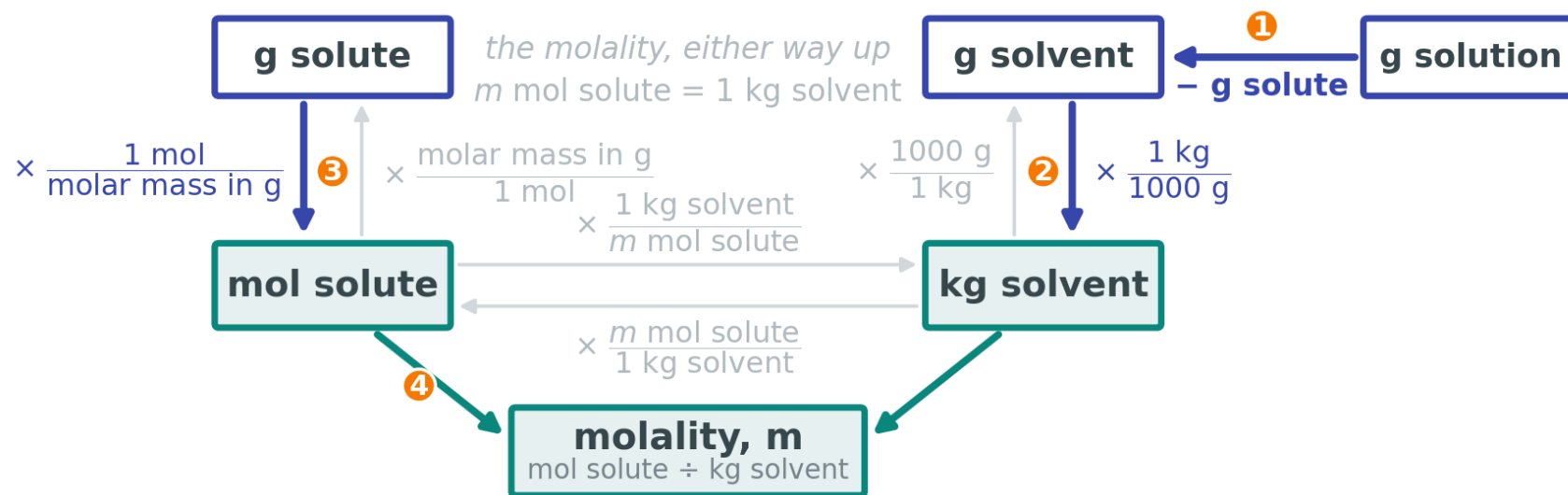
A used the solution's mass: $0.500 \div 0.2873 = 1.74$. B stopped at moles. C added the KCl to a total that already held it: $0.500 \div 0.3246 = 1.54$. D divided by grams: $0.500 \div 250.0 = 2.00 \times 10^{-3}$.

 **Sense check:** A quarter kilogram of water carries 0.500 mol, so a full kilogram carries four times that: 2.00 mol ✓

✓ Practice 5: the route on the map

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 37.3 g KCl · 287.3 g solution · found: 2.00 mol/kg



🧠 **Sense check:** Four moves: the KCl leaves the total to give the water, the water becomes kilograms, the KCl becomes moles, then moles ÷ kilograms. ✓



Practice 6

solubility of NH_4Cl at $30\text{ }^\circ\text{C}$: 41.4 g per 100 g of water

molar mass: NH_4Cl 53.49 g/mol

A saturated NH_4Cl solution at $30\text{ }^\circ\text{C}$ holds **41.4 g of NH_4Cl** in every **100. g of water**. What is its molality?

- A. 0.774
- B. 5.47
- C. 7.74
- D. 0.129

✓ Practice 6 answer: C

$$m = \text{mol solute} \div \text{kg solvent}$$

given: 41.4 g NH₄Cl (53.49 g/mol) per 100. g water · saturated at 30 °C · wanted: m

$$41.4 \text{ g NH}_4\text{Cl} \times \frac{1 \text{ mol NH}_4\text{Cl}}{53.49 \text{ g NH}_4\text{Cl}} = 0.774 \text{ mol} \cdot 100. \text{ g water} = 0.100 \text{ kg water}$$

$$m = \frac{0.774 \text{ mol NH}_4\text{Cl}}{0.100 \text{ kg water}} = \mathbf{7.74 \text{ mol/kg} \cdot \text{answer C}}$$

A stopped at moles: 0.774 mol sits in 100. g of water, not in a kilogram. B used the solution's mass: $0.774 \div 0.1414 \text{ kg} = 5.47$, but a solubility counts 100 g of water, not of solution. D flipped the fraction: $0.100 \div 0.774 = 0.129$ counts kilograms per mole.

 **Sense check:** A tenth of a kilogram of water holds 0.774 mol, so a full kilogram holds ten times that: 7.74 mol ✓

 Practice 7

aqueous potassium bromide (KBr)

molar mass: KBr 119.00 g/mol

A **0.400 m** KBr solution is made by dissolving **23.8 g of KBr** in water. What is the total mass of the solution, in grams?

- A. 500.
- B. 524
- C. 104
- D. 476

✓ Practice 7 answer: B

0.400 m KBr: 0.400 mol KBr = 1 kg water · solution = solute + solvent
given: 23.8 g KBr (119.00 g/mol) · 0.400 m · wanted: g solution · three conversion factors

$$23.8 \text{ g KBr} \times \frac{1 \text{ mol KBr}}{119.00 \text{ g KBr}} \times \frac{1 \text{ kg water}}{0.400 \text{ mol KBr}} \times \frac{1000 \text{ g water}}{1 \text{ kg water}} = 500. \text{ g water}$$

$$500. \text{ g water} + 23.8 \text{ g KBr} = \mathbf{524 \text{ g solution} \cdot \text{answer B}}$$

A stopped at the water: 500. g is the solvent alone. C flipped the molality factor: $0.200 \times 0.400 = 0.0800$ kg of water, plus 23.8 g = 104. D subtracted the KBr: $500. - 23.8 = 476$, as if 500. g were the whole solution.

 **Sense check:** 0.200 mol at 0.400 mol per kilogram needs half a kilogram of water. The KBr adds 23.8 g: 524 g ✓

Check yourself

1. A solution is prepared from grams of solute and grams of water. Which mass belongs in the denominator of molality, the solvent's or the whole solution's, and in what unit?
2. A 2.0 m solution is warmed from 20 °C to 60 °C. Is it still 2.0 m? Explain with mass and volume.

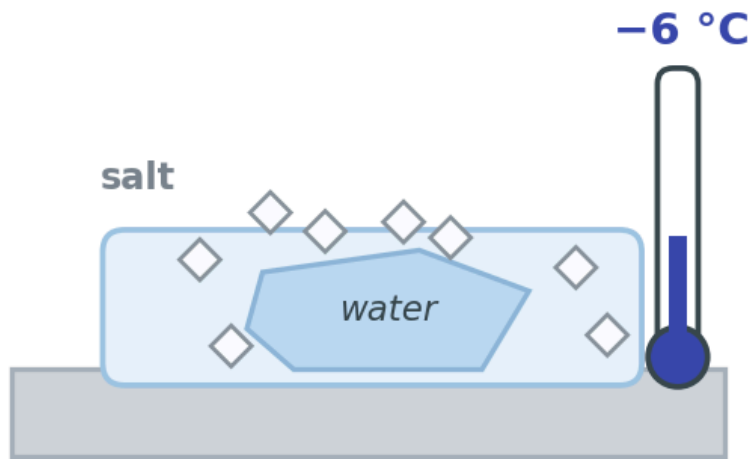
Molarity counts moles per liter of solution; molality counts moles per kilogram of solvent. Molality returns in freezing-point and boiling-point changes, where the shift grows with the solute's moles per kilogram of solvent.

8 · Freezing Point & Boiling Point

Count the dissolved particles with the van't Hoff factor i , then use $\Delta T_f = i \cdot K_f \cdot m$ and $\Delta T_b = i \cdot K_b \cdot m$ to find how far a solute lowers the freezing point and raises the boiling point of water.

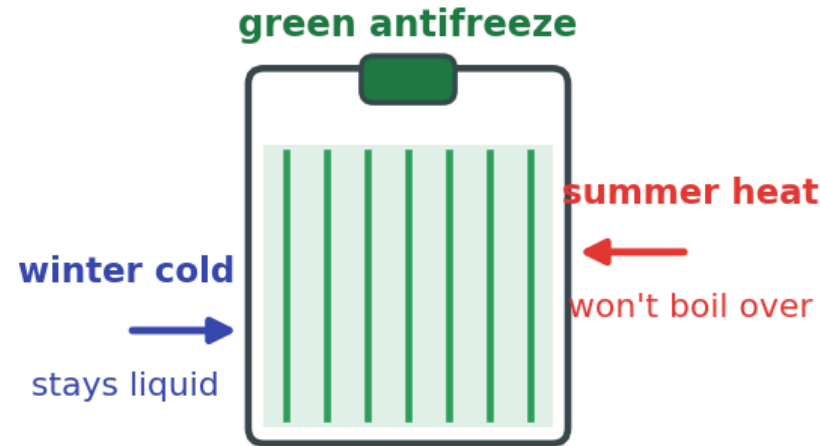
Salt on ice, antifreeze in a radiator

road salt on an icy walk



ice melts to water at $-6\text{ }^{\circ}\text{C}$

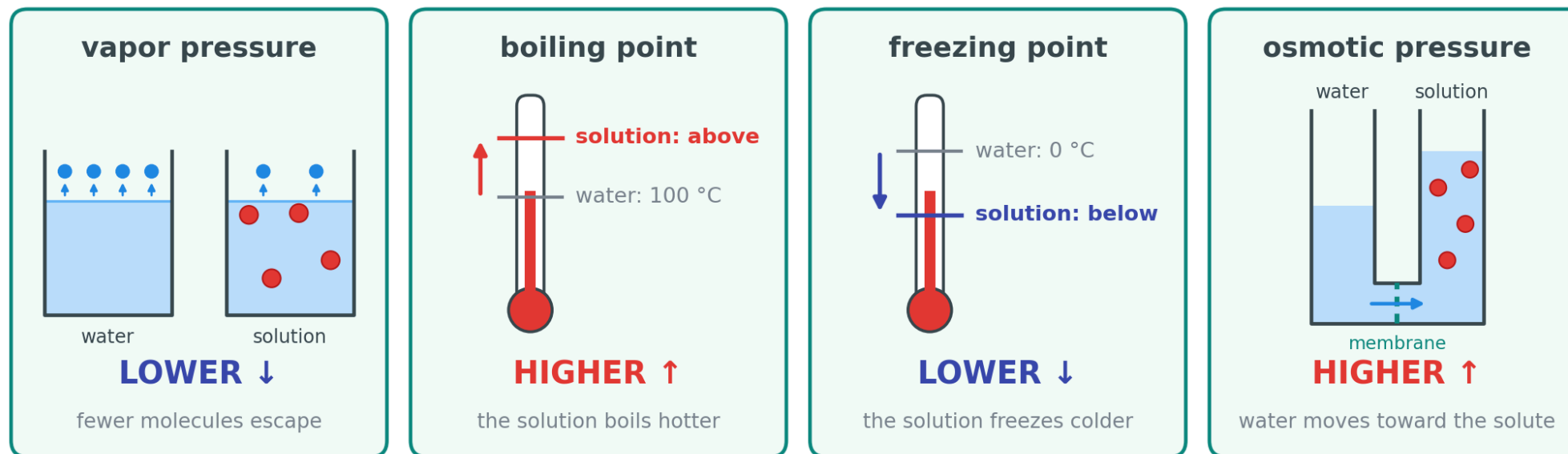
antifreeze in a radiator



liquid when cold, no boil-over when hot

Scatter salt on a frozen walk and the ice melts below $0\text{ }^{\circ}\text{C}$. Antifreeze keeps a radiator from freezing in winter and boiling over in summer.

Four properties that count particles

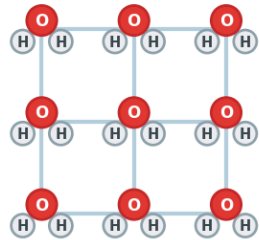


each change grows with the NUMBER of dissolved particles, not their identity

A colligative property depends only on how many solute particles dissolve, not which ones. Osmotic pressure is the push needed to stop water from crossing a membrane into a solution.

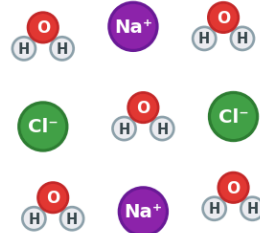
Freezing lower, boiling higher

dissolved particles get in the way of freezing



pure water

molecules lock into an ordered solid



salt water

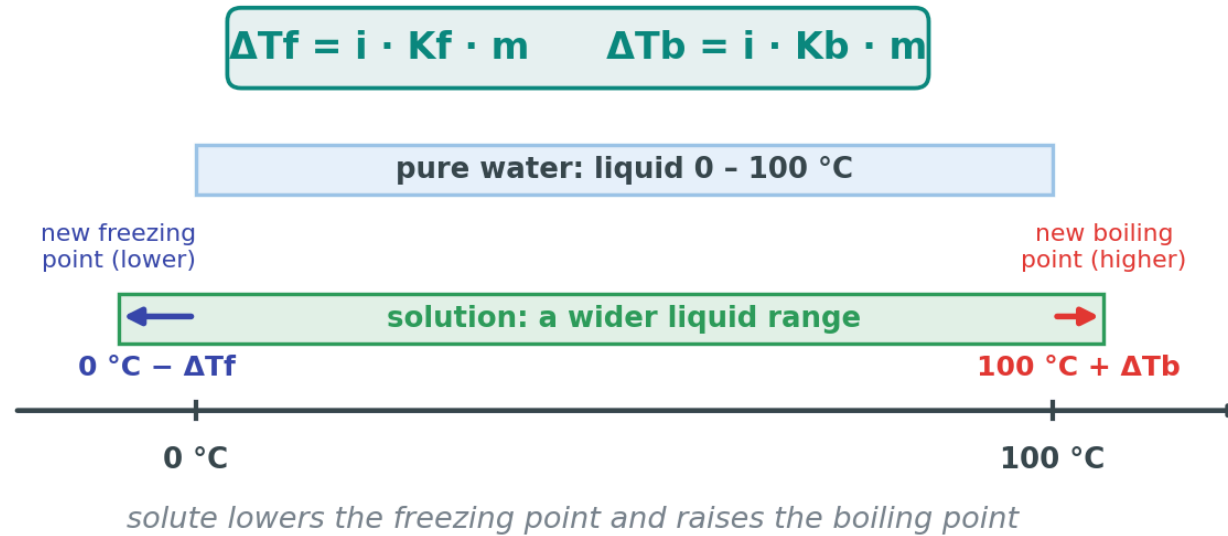
particles block the lattice — must go colder to freeze

only the **NUMBER** of dissolved particles matters, not what they are

Freezing builds an ordered lattice; any particle blocks it, so the solution freezes colder. Solute also crowds the liquid's surface; fewer molecules escape, so the liquid must run hotter to boil. Number matters, not identity.

memory hook: a solute stretches the liquid range both ways
freezes below the pure solvent · boils above it · more particles, more stretch

The size of the shift



water: $K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$ $K_b = 0.512 \text{ }^\circ\text{C}\cdot\text{kg/mol}$

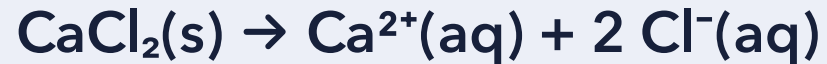
$\Delta T_f = i \cdot K_f \cdot m$ (down) · $\Delta T_b = i \cdot K_b \cdot m$ (up)

Each equation multiplies the particle count i , the solvent constant K , and the molality m . K_f sets the freezing shift, K_b the boiling shift. A bigger i or higher m shifts more.

Recap: what a solute becomes in water



strong electrolyte · 1 + 1 = 2 particles



strong electrolyte · 1 + 2 = 3 particles

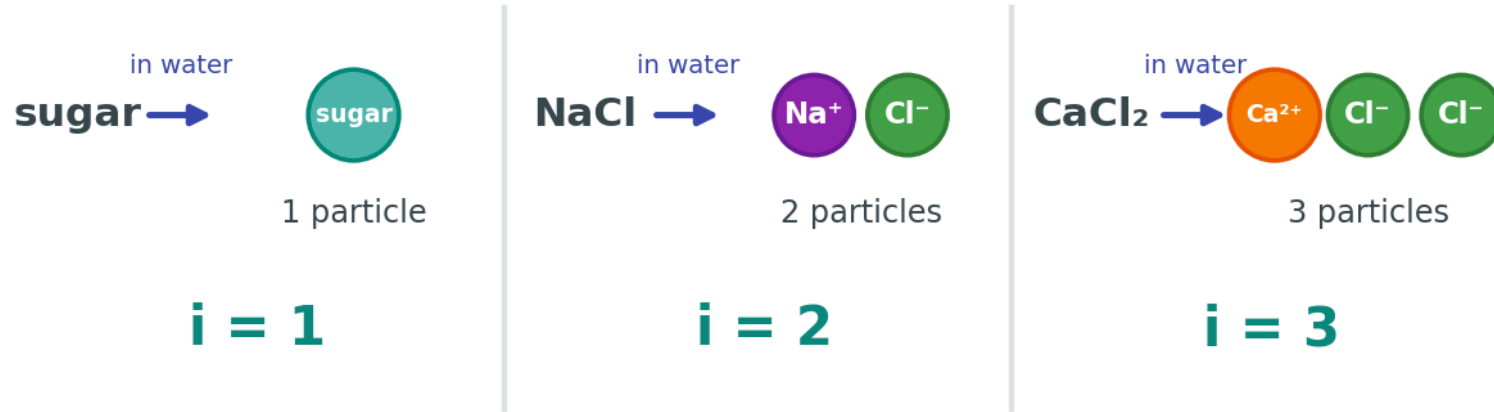


sugar · nonelectrolyte · dissolves whole · 1 particle

A strong electrolyte dissociates into separate ions. A nonelectrolyte dissolves as whole molecules. Dissociation sorts solutes into these classes. Freezing and boiling shifts respond only to the particle count.

The van't Hoff factor i

the van't Hoff factor i = particles per formula unit



more particles → a bigger freezing and boiling shift

i counts how many particles each formula unit releases. A molecular solute stays whole, so $i = 1$. NaCl gives two ions, CaCl₂ three. More particles mean a bigger shift.

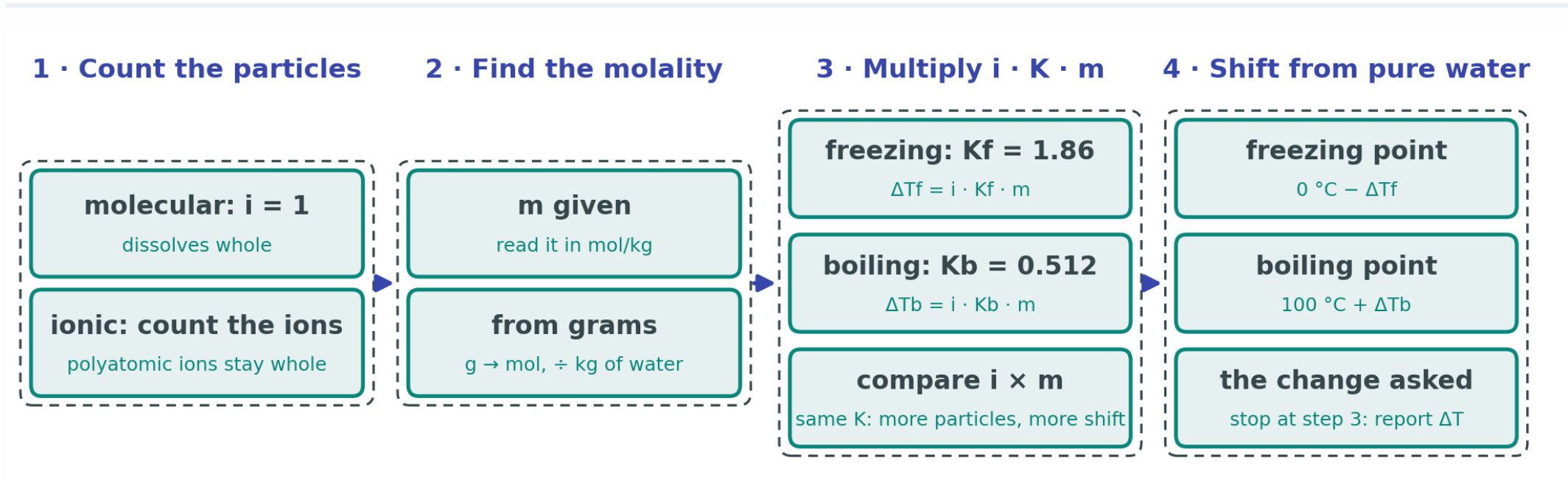
memory hook: i = the pieces the formula splits into
glucose 1 · NaCl 2 · CaCl₂ 3 (one Ca²⁺, two Cl⁻) · Na₂SO₄ 3 (two Na⁺, one SO₄²⁻)



The method

1. **Count the particles.** Read i from the formula.
2. **Find the molality.** $\text{mol solute} \div \text{kg SOLVENT}$.
3. **Multiply $i \cdot K \cdot m$.** K_f 1.86 freezing, K_b 0.512 boiling.
4. **Shift from pure water.** $0 - \Delta T_f$, or $100 + \Delta T_b$.

The route to a freezing or boiling point



A problem lights one path through the four columns. A ranking question compares $i \times m$. A question that asks for ΔT stops after step 3.



Guided example: which boils higher?

$$\Delta T_b = i \cdot K_b \cdot m$$

given: 0.800 m glucose, $C_6H_{12}O_6$ · 0.800 m KBr · both in water · wanted: the solution with the higher boiling point

Two pots of water each hold **0.800 m** of solute: glucose in one, potassium bromide in the other. Which boils at the higher temperature?

Three moves: count the particles, multiply by the molality, compare.

✓ Guided example: solution

$$\Delta T_b = i \cdot K_b \cdot m$$

0.800 m glucose · 0.800 m KBr · same water, same $K_b = 0.512 \text{ }^\circ\text{C}\cdot\text{kg/mol}$

Step 1 · Count the particles

glucose: molecular, dissolves whole · KBr: $\text{K}^+ + \text{Br}^-$

glucose $i = 1$ · KBr $i = 2$

Compare the particle counts

glucose: $1 \times 0.800 \text{ m} = 0.800$ · KBr: $2 \times 0.800 \text{ m} = 1.60$

moles of dissolved particles per kg of water

 **Sense check:** Check with the equation: glucose $1 \times 0.512 \times 0.800 = 0.410 \text{ }^\circ\text{C}$, KBr $2 \times 0.512 \times 0.800 = 0.819 \text{ }^\circ\text{C}$.
KBr boils higher: $100.82 \text{ }^\circ\text{C}$ against $100.41 \text{ }^\circ\text{C}$. ✓

✓ Guided example: the route on the map

KBr: $2 \times 0.800 = 1.60$ · glucose: $1 \times 0.800 = 0.800$
 given: 0.800 m each · found: KBr boils higher

1 · Count the particles

molecular: $i = 1$
 glucose: $i = 1$

ionic: count the ions
 KBr: $i = 2$

2 · Find the molality

m given
 0.800 m each

from grams
 $\text{g} \rightarrow \text{mol}, \div \text{kg of water}$

3 · Multiply $i \cdot K \cdot m$

freezing: $K_f = 1.86$
 $\Delta T_f = i \cdot K_f \cdot m$

boiling: $K_b = 0.512$
 $\Delta T_b = i \cdot K_b \cdot m$

compare $i \times m$
 1.60 vs 0.800

4 · Shift from pure water

freezing point
 $0^\circ\text{C} - \Delta T_f$

boiling point
 KBr boils higher

the change asked
 stop at step 3: report ΔT

🧠 **Sense check:** A ranking needs only $i \times m$. The same K cancels out of the comparison. ✓



Worked example 1: freezing point of antifreeze

$$\Delta T_f = i \cdot K_f \cdot m$$

given: 1.50 molal ethylene glycol · $K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$ · wanted: ΔT_f

A radiator holds **1.50 molal ethylene glycol** in water. Ethylene glycol is molecular: it dissolves without splitting into ions.

Count the particles, then multiply.

✓ Worked example 1: solution

$$\Delta T_f = i \cdot K_f \cdot m$$

given: 1.50 molal ethylene glycol · $K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$ · wanted: ΔT_f

Step 1 · Count the particles

Step 2 · Find the molality

Step 3 · Multiply $i \cdot K \cdot m$

$m = 1.50 \text{ mol/kg}$,

given.

$$\Delta T_f = 1 \times \frac{1.86 \text{ }^\circ\text{C}\cdot\cancel{\text{kg}}}{\cancel{\text{mol}}} \times \frac{1.50 \cancel{\text{mol}}}{\cancel{\text{kg}}} = 2.79 \text{ }^\circ\text{C}$$

 **Sense check:** A molecular solute still shifts the point. The freezing point drops to $0 - 2.79 = -2.79 \text{ }^\circ\text{C}$. Freezing point goes DOWN. ✓

✓ Worked example 1: the route on the map

$$\Delta T_f = 1 \times 1.86 \times 1.50 = 2.79 \text{ }^{\circ}\text{C}$$

ethylene glycol, molecular: $i = 1$ · m given · the change is asked, so stop at ΔT_f

1 · Count the particles

molecular: $i = 1$
ethylene glycol: $i = 1$

ionic: count the ions
polyatomic ions stay whole

2 · Find the molality

m given
1.50 m

from grams
g → mol, ÷ kg of water

3 · Multiply $i \cdot K \cdot m$

freezing: $K_f = 1.86$
 $1 \times 1.86 \times 1.50$

boiling: $K_b = 0.512$
 $\Delta T_b = i \cdot K_b \cdot m$

compare $i \times m$
same K : more particles, more shift

4 · Shift from pure water

freezing point
 $0 \text{ }^{\circ}\text{C} - \Delta T_f$

boiling point
 $100 \text{ }^{\circ}\text{C} + \Delta T_b$

the change asked
 $\Delta T_f = 2.79 \text{ }^{\circ}\text{C}$

🧠 **Sense check:** One multiplication. A molecular solute keeps $i = 1$, and m was given. ✓



Worked example 2: freezing point of a salt solution

$$\Delta T_f = i \cdot K_f \cdot m$$

given: 1.20 molal NaCl · $K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$ · wanted: ΔT_f

A **1.20 molal NaCl** solution is spread on an icy road. A common first attempt multiplies K_f by m and stops.

Count the particles, then multiply.

✓ Worked example 2: solution

$$\Delta T_f = i \cdot K_f \cdot m$$

given: 1.20 molal NaCl · $K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$ · wanted: ΔT_f

A common first attempt

$$\Delta T_f = \frac{1.86 \text{ }^\circ\text{C}\cdot\text{kg}}{\text{mol}} \times \frac{1.20 \text{ mol}}{\text{kg}} = 2.23 \text{ }^\circ\text{C} \times$$

Step 1 · Count the particles

Step 2 · Find the molality

Step 3 · Multiply $i \cdot K \cdot m$

$$\Delta T_f = 2 \times \frac{1.86 \text{ }^\circ\text{C}\cdot\text{kg}}{\text{mol}} \times \frac{1.20 \text{ mol}}{\text{kg}} = 4.46 \text{ }^\circ\text{C}$$

🧠 **Sense check:** Two ions double the depression to $4.46 \text{ }^\circ\text{C}$. Freezing point goes DOWN, to $-4.46 \text{ }^\circ\text{C}$. ✓

✓ Worked example 2: the route on the map

$$\Delta T_f = 2 \times 1.86 \times 1.20 = 4.46 \text{ }^{\circ}\text{C}$$

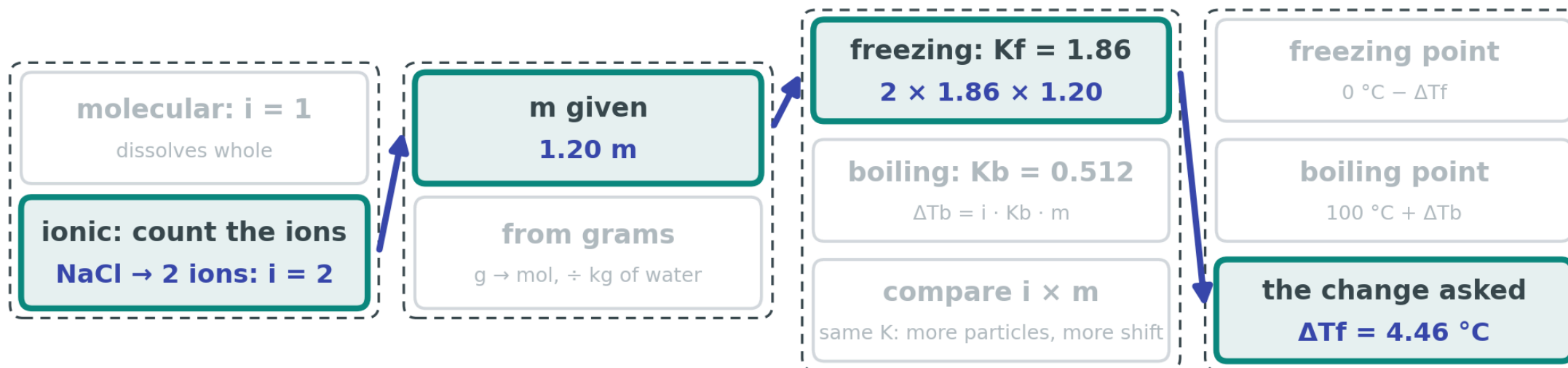
NaCl gives $\text{Na}^+ + \text{Cl}^-$: $i = 2$ · m given · stop at ΔT_f

1 · Count the particles

2 · Find the molality

3 · Multiply $i \cdot K \cdot m$

4 · Shift from pure water



🧠 **Sense check:** The same route as example 1. Only step 1 changed: two ions, twice the shift. ✓

Take-home: count the particles

1.20 molal NaCl · K_f = 1.86

forgot i (i = 1): 1.86 × 1.20 = 2.23 °C × · counted i = 2: 2 × 1.86 × 1.20 = 4.46 °C ✓

An ionic solute breaks into ions, and each ion counts. Skip i and the depression comes out too small. That is the usual mistake. Read i from the formula before multiplying.

 Practice 1

$$\Delta T_f = i \cdot K_f \cdot m$$

$K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$ for every solution below

Three solutions are each **0.40 m** in water. Which one has the lowest freezing point?

- A. sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$
- B. calcium chloride, CaCl_2
- C. potassium nitrate, KNO_3
- D. all three freeze at the same temperature

✓ Practice 1 answer: B

$i \times m$: sucrose $1 \times 0.40 = 0.40$ · KNO_3 $2 \times 0.40 = 0.80$ · CaCl_2 $3 \times 0.40 = 1.20$

CaCl_2 gives $\text{Ca}^{2+} + 2 \text{Cl}^-$: the most particles · the lowest freezing point · answer B

A judged by the size of the molecule: sucrose's 45 atoms stay together as one particle, so $i = 1$. C split the nitrate into atoms, $1 + 1 + 3 = 5$ pieces; a polyatomic ion stays whole, so KNO_3 gives K^+ and NO_3^- , $i = 2$. D ignored i : equal molality is not equal particle count; 0.40 m of CaCl_2 is 1.20 mol/kg of particles.

🧠 **Sense check:** 1.20 mol/kg of particles is the most, so CaCl_2 freezes lowest, near $0 - 3 \times 1.86 \times 0.40 = -2.2^\circ\text{C}$. ✓

✓ Practice 1: the route on the map

CaCl_2 1.20 > KNO_3 0.80 > sucrose 0.40
given: 0.40 m each · found: CaCl_2 freezes lowest

1 · Count the particles

molecular: $i = 1$
sucrose: $i = 1$

ionic: count the ions
 CaCl_2 3 · KNO_3 2

2 · Find the molality

m given
0.40 m each

from grams
g → mol, ÷ kg of water

3 · Multiply $i \cdot K \cdot m$

freezing: $K_f = 1.86$
 $\Delta T_f = i \cdot K_f \cdot m$

boiling: $K_b = 0.512$
 $\Delta T_b = i \cdot K_b \cdot m$

compare $i \times m$
1.20 > 0.80 > 0.40

4 · Shift from pure water

freezing point
lowest: CaCl_2

boiling point
100 °C + ΔT_b

the change asked
stop at step 3: report ΔT

🧠 **Sense check:** The guided example's route: count, multiply by m, compare. ✓

 Practice 2

$$\Delta T_f = i \cdot K_f \cdot m$$

given: 0.900 molal KCl · $K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$ · wanted: ΔT_f

A **0.900 molal KCl** solution is used as a de-icer. What is its freezing-point depression, in $^\circ\text{C}$?

- A. 1.67
- B. 0.922
- C. 3.35
- D. 1.80

✓ Practice 2 answer: C

$$\Delta T_f = i \cdot K_f \cdot m$$

$$\text{given: } 0.900 \text{ molal KCl} \cdot i = 2 \cdot K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$$

$$\Delta T_f = 2 \times \frac{1.86 \text{ }^\circ\text{C}\cdot\text{kg}}{\text{mol}} \times \frac{0.900 \text{ mol}}{\text{kg}} = 3.35 \text{ }^\circ\text{C} \cdot \text{answer C}$$

A forgot i : $1.86 \times 0.900 = 1.67 \text{ }^\circ\text{C}$ uses $i = 1$, but KCl gives two ions. B used K_b : $2 \times 0.512 \times 0.900 = 0.922 \text{ }^\circ\text{C}$ is the boiling constant, not the freezing one. D stopped at $i \times m$: $2 \times 0.900 = 1.80 \text{ mol/kg}$ counts the dissolved particles but never multiplies by K_f .

 **Sense check:** Two ions and $K_f = 1.86$ give a $3.35 \text{ }^\circ\text{C}$ depression. The freezing point drops to $-3.35 \text{ }^\circ\text{C}$. Freezing point goes DOWN. ✓

✓ Practice 2: the route on the map

$$\Delta T_f = 2 \times 1.86 \times 0.900 = 3.35 \text{ }^{\circ}\text{C}$$

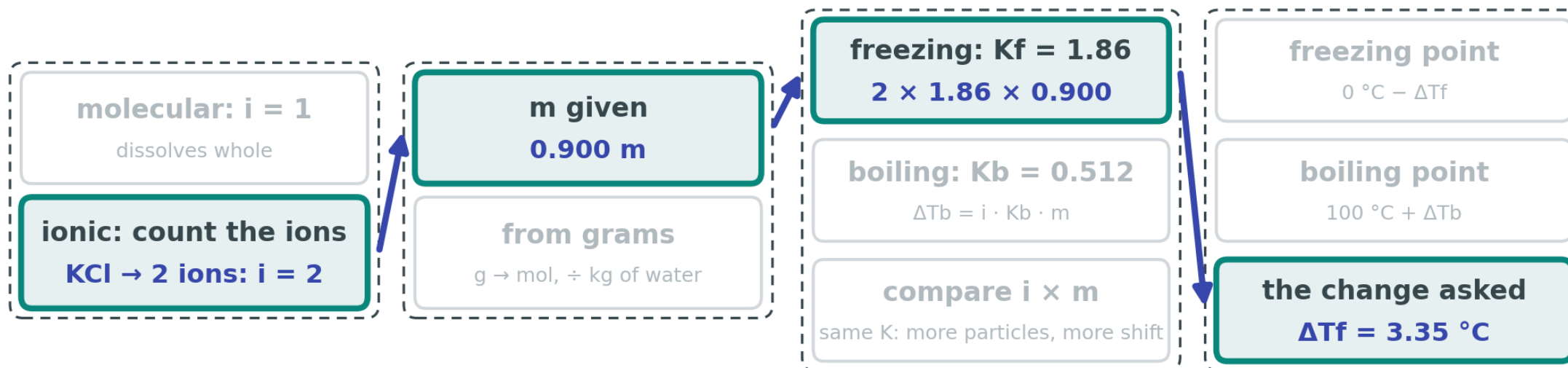
KCl gives $\text{K}^+ + \text{Cl}^-$: $i = 2$ · m given · the change is asked: stop at ΔT_f

1 · Count the particles

2 · Find the molality

3 · Multiply $i \cdot K \cdot m$

4 · Shift from pure water



🧠 **Sense check:** The route of worked example 2, with KCl's two ions. ✓



Worked example 3: new freezing point from grams

$$\text{new freezing point} = 0\text{ }^{\circ}\text{C} - \Delta T_f$$

given: 41.6 g CaCl_2 (110.98 g/mol) · 500. g water · $K_f = 1.86\text{ }^{\circ}\text{C}\cdot\text{kg/mol}$ · wanted: the new freezing point

A road brine is mixed from **41.6 g of CaCl_2** and **500. g of water**. A common first attempt reports the depression itself as the freezing point.

Count the particles, find the molality, then shift from $0\text{ }^{\circ}\text{C}$.

✓ Worked example 3: solution

new freezing point = $0\text{ }^{\circ}\text{C} - \Delta T_f$

given: 41.6 g CaCl_2 (110.98 g/mol) · 500. g water · $K_f = 1.86\text{ }^{\circ}\text{C}\cdot\text{kg/mol}$ · wanted: the new freezing point

Step 1 · Count the particles

Step 2 · Find the molality

$$41.6 \text{ g } \text{CaCl}_2 \times \frac{1 \text{ mol } \text{CaCl}_2}{110.98 \text{ g } \text{CaCl}_2} = 0.3748 \text{ mol, then } \frac{0.3748 \text{ mol } \text{CaCl}_2}{0.500 \text{ kg water}} = \mathbf{0.7496 \text{ mol/kg}}$$

The denominator is the water alone, 0.500 kg, not the 0.5416 kg of solution.

🧠 **Sense check:** $i = 3$ and $m = 0.7496 \text{ mol/kg}$ carry into the depression. ✓

✓ Worked example 3: the new freezing point

$$\text{new freezing point} = 0\text{ }^{\circ}\text{C} - \Delta T_f$$

$$41.6\text{ g CaCl}_2 \text{ in } 500.\text{ g water} \cdot i = 3 \cdot m = 0.7496\text{ mol/kg} \cdot K_f = 1.86\text{ }^{\circ}\text{C}\cdot\text{kg/mol}$$

Step 3 • Multiply $i \cdot K \cdot m$

$$\Delta T_f = 3 \times \frac{1.86\text{ }^{\circ}\text{C}\cdot\cancel{\text{kg}}}{\cancel{\text{mol}}} \times \frac{0.7496\cancel{\text{ mol}}}{\cancel{\text{kg}}} = 4.18\text{ }^{\circ}\text{C}$$

Step 4 • Shift from pure water

$$\text{new freezing point} = 0\text{ }^{\circ}\text{C} - 4.18\text{ }^{\circ}\text{C} = -4.18\text{ }^{\circ}\text{C}$$

 **Sense check:** The depression is $4.18\text{ }^{\circ}\text{C}$, but the freezing point is $0 - 4.18 = -4.18\text{ }^{\circ}\text{C}$, below pure water's, not $+4.18$. Freezing point goes DOWN. ✓

✓ Worked example 3: the route on the map

$$41.6 \text{ g CaCl}_2 \rightarrow 0.7496 \text{ m} \rightarrow \Delta T_f = 4.18 \text{ }^\circ\text{C} \rightarrow -4.18 \text{ }^\circ\text{C}$$

$i = 3$ · molality from grams, ÷ 0.500 kg of water · the freezing point is asked: shift from 0 °C

1 · Count the particles

molecular: $i = 1$

dissolves whole

ionic: count the ions

$\text{CaCl}_2 \rightarrow 3 \text{ ions: } i = 3$

2 · Find the molality

m given

read it in mol/kg

from grams

$41.6 \text{ g} \rightarrow 0.7496 \text{ m}$

3 · Multiply $i \cdot K \cdot m$

freezing: $K_f = 1.86$

$\Delta T_f = 4.18 \text{ }^\circ\text{C}$

boiling: $K_b = 0.512$

$\Delta T_b = i \cdot K_b \cdot m$

compare $i \times m$

same K : more particles, more shift

4 · Shift from pure water

freezing point

$0 - 4.18 = -4.18 \text{ }^\circ\text{C}$

boiling point

$100 \text{ }^\circ\text{C} + \Delta T_b$

the change asked

stop at step 3: report ΔT

🧠 **Sense check:** All four steps light up. Grams add the molality step. ✓

Your turn: new boiling point of a salt solution

$$\text{new boiling point} = 100\text{ }^{\circ}\text{C} + \Delta T_b$$

given: 1.00 molal MgCl_2 · $K_b = 0.512\text{ }^{\circ}\text{C}\cdot\text{kg/mol}$ · wanted: the new boiling point

MgCl_2 dissolves into one Mg^{2+} and two Cl^- . Fill in i and K_b , find ΔT_b , then add.

$$\Delta T_b = \text{---} \times \text{---} \times 1.00 = \text{---}^{\circ}\text{C}$$

$$\Delta T_b = 3 \times \frac{0.512\text{ }^{\circ}\text{C}\cdot\cancel{\text{kg}}}{\cancel{\text{mol}}} \times \frac{1.00\cancel{\text{ mol}}}{\cancel{\text{kg}}} = 1.54\text{ }^{\circ}\text{C}$$

$$\text{new boiling point} = 100\text{ }^{\circ}\text{C} + 1.54\text{ }^{\circ}\text{C} = 101.54\text{ }^{\circ}\text{C}$$

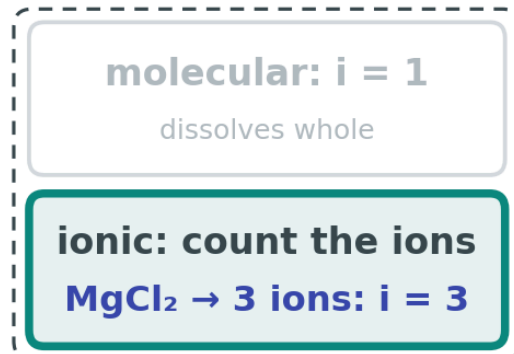
 **Sense check:** Three ions and $K_b = 0.512$ raise the boiling point by $1.54\text{ }^{\circ}\text{C}$, to $101.54\text{ }^{\circ}\text{C}$. Boiling point goes UP. ✓

✓ Your turn: the route on the map

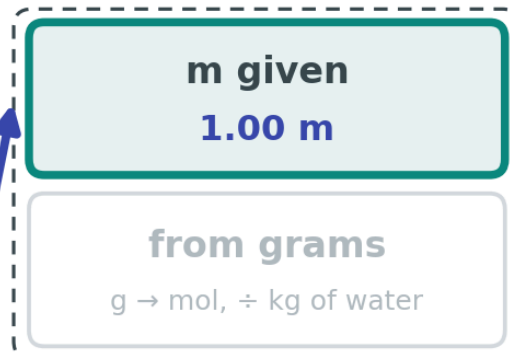
$$\Delta T_b = 3 \times 0.512 \times 1.00 = 1.54 \text{ }^{\circ}\text{C} \rightarrow 101.54 \text{ }^{\circ}\text{C}$$

MgCl_2 : $i = 3 \cdot m$ given $\cdot$ boiling uses K_b , then $100 \text{ }^{\circ}\text{C} + \Delta T_b$

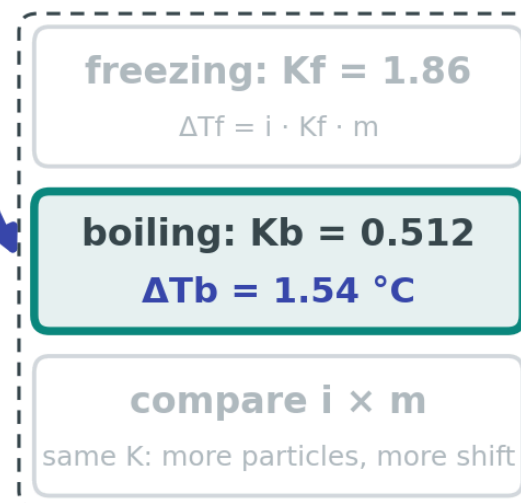
1 · Count the particles



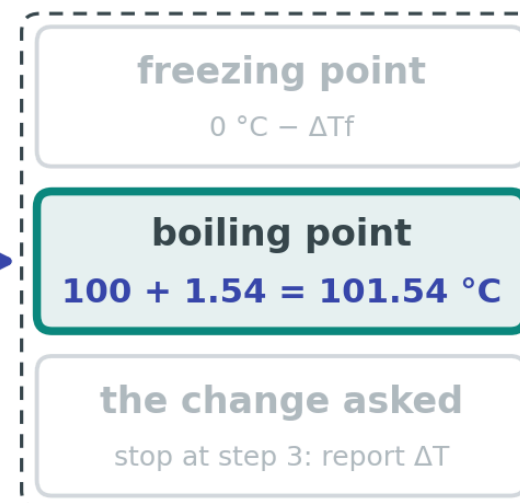
2 · Find the molality



3 · Multiply $i \cdot K \cdot m$



4 · Shift from pure water



🧠 **Sense check:** Boiling swaps K_f for K_b and the minus for a plus. ✓

⚠ Where this goes wrong

$$\Delta T_f = i \cdot K_f \cdot m$$

$$1.20 \text{ molal NaCl} \cdot i = 2 \cdot K_f = 1.86 \cdot \text{correct } \Delta T_f = 4.46 \text{ }^\circ\text{C}$$

Forgetting the particle count. Using $i = 1$ gives $1.86 \times 1.20 = 2.23 \text{ }^\circ\text{C}$. NaCl releases two ions, so $i = 2$ and $\Delta T_f = 4.46 \text{ }^\circ\text{C}$, twice as much.

Using K_b for a freezing problem. $2 \times 0.512 \times 1.20 = 1.23 \text{ }^\circ\text{C}$ uses the boiling constant. Freezing-point depression uses $K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$.

Dividing by the solution's mass. $0.3748 \text{ mol CaCl}_2 \div 0.5416 \text{ kg of solution} = 0.692 \text{ m}$ counts the salt as solvent. Molality divides by the water alone: $0.3748 \div 0.500 = 0.7496 \text{ m}$.

Reporting the depression as the temperature. A $4.18 \text{ }^\circ\text{C}$ depression is not a $4.18 \text{ }^\circ\text{C}$ freezing point. The new point is $0 - 4.18 = -4.18 \text{ }^\circ\text{C}$.



Practice 3

$$\text{new freezing point} = 0\text{ }^{\circ}\text{C} - \Delta T_f$$

$$K_f = 1.86\text{ }^{\circ}\text{C}\cdot\text{kg/mol}$$

A lawn-fertilizer solution is **1.30 m** ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$, in water. At what temperature, in $^{\circ}\text{C}$, does it freeze?

- A. -7.25
- B. -4.84
- C. -2.42
- D. 7.25
- E. -2.00

✓ Practice 3 answer: A



given: $1.30 \text{ m} \cdot K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$ · each ammonium ion stays whole

$$\Delta T_f = 3 \times 1.86 \times 1.30 = 7.25 \text{ }^\circ\text{C} \rightarrow 0 \text{ }^\circ\text{C} - 7.25 \text{ }^\circ\text{C} = \mathbf{-7.25 \text{ }^\circ\text{C} \cdot \text{answer A}}$$

B counted one NH_4^+ and one SO_4^{2-} : $i = 2$ gives $-4.84 \text{ }^\circ\text{C}$, but the 2 outside the parentheses means two ammonium ions. C forgot i : $0 - 1.86 \times 1.30 = -2.42 \text{ }^\circ\text{C}$. D stopped at the depression and dropped the sign: $+7.25 \text{ }^\circ\text{C}$ is above $0 \text{ }^\circ\text{C}$. E used K_b : $0 - 3 \times 0.512 \times 1.30 = -2.00 \text{ }^\circ\text{C}$.

 **Sense check:** Three ions per formula unit push the freezing point to $-7.25 \text{ }^\circ\text{C}$, below pure water's $0 \text{ }^\circ\text{C}$. Freezing point goes DOWN. ✓

✓ Practice 3: the route on the map

$$\Delta T_f = 3 \times 1.86 \times 1.30 = 7.25 \text{ }^{\circ}\text{C} \rightarrow -7.25 \text{ }^{\circ}\text{C}$$

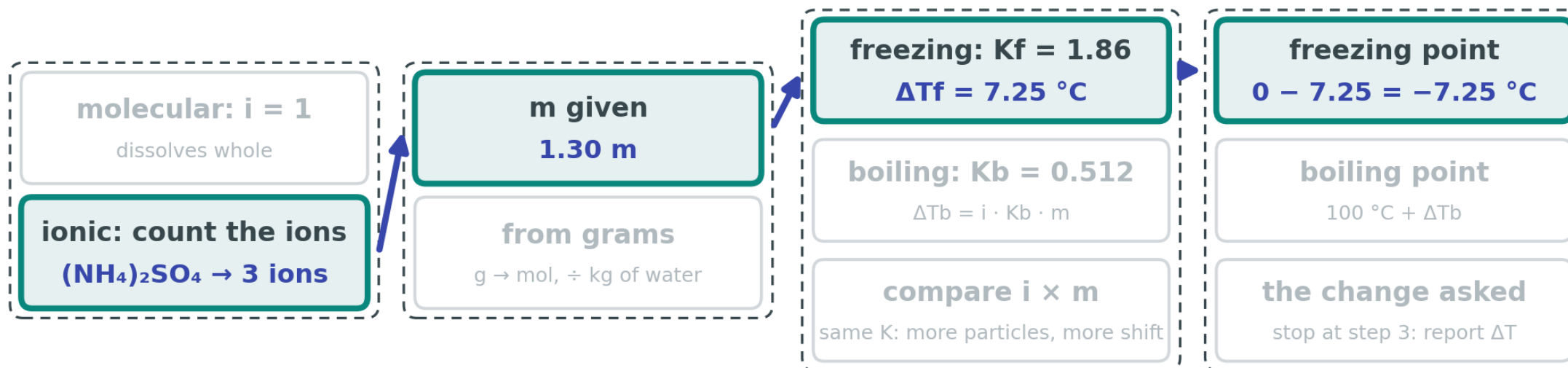
$(\text{NH}_4)_2\text{SO}_4$: $i = 3 \cdot m$ given $\cdot$ the freezing point is asked: shift from $0 \text{ }^{\circ}\text{C}$

1 · Count the particles

2 · Find the molality

3 · Multiply $i \cdot K \cdot m$

4 · Shift from pure water



🧠 **Sense check:** Step 1 carried the weight: the subscript outside the parentheses counts ions. ✓



Practice 4

$$\text{new freezing point} = 0\text{ }^{\circ}\text{C} - \Delta T_f$$

$$\text{molar mass BaCl}_2 \text{ } 208.23 \text{ g/mol} \cdot K_f = 1.86\text{ }^{\circ}\text{C}\cdot\text{kg/mol}$$

A solution is made by dissolving **36.4 g of BaCl₂** in **250. g of water**. What is its freezing point, in $^{\circ}\text{C}$?

- A. -1.30
- B. 3.90
- C. -3.41
- D. -3.90

✓ Practice 4 answer: D

$$\text{new freezing point} = 0\text{ }^{\circ}\text{C} - \Delta T_f$$

$$\text{given: } 36.4\text{ g BaCl}_2 (208.23\text{ g/mol}) \cdot 250.\text{ g water} \cdot i = 3 \cdot K_f = 1.86\text{ }^{\circ}\text{C}\cdot\text{kg/mol}$$

$$36.4\text{ g} \div 208.23\text{ g/mol} = 0.1748\text{ mol BaCl}_2 \rightarrow \frac{0.1748\text{ mol}}{0.250\text{ kg water}} = 0.699\text{ mol/kg}$$

$$\Delta T_f = 3 \times 1.86 \times 0.699 = 3.90\text{ }^{\circ}\text{C} \rightarrow 0\text{ }^{\circ}\text{C} - 3.90\text{ }^{\circ}\text{C} = \mathbf{-3.90\text{ }^{\circ}\text{C} \cdot \text{answer D}}$$

A forgot i : $0 - 1.86 \times 0.699 = -1.30\text{ }^{\circ}\text{C}$, but BaCl_2 gives three ions. B stopped at the depression and dropped the sign: $+3.90\text{ }^{\circ}\text{C}$ is above $0\text{ }^{\circ}\text{C}$. C divided by the solution's mass: $0.1748 \div 0.2864\text{ kg} = 0.610\text{ m}$, giving $-3.41\text{ }^{\circ}\text{C}$.

 **Sense check:** Freezing point goes DOWN: $-3.90\text{ }^{\circ}\text{C}$, below pure water's $0\text{ }^{\circ}\text{C}$. ✓

✓ Practice 4: the route on the map

$$36.4 \text{ g BaCl}_2 \rightarrow 0.699 \text{ m} \rightarrow \Delta T_f = 3.90 \text{ }^\circ\text{C} \rightarrow -3.90 \text{ }^\circ\text{C}$$

BaCl_2 : $i = 3$ · molality from grams, $\div 0.250 \text{ kg of water}$ · shift from $0 \text{ }^\circ\text{C}$

1 · Count the particles

molecular: $i = 1$

dissolves whole

ionic: count the ions

$\text{BaCl}_2 \rightarrow 3 \text{ ions: } i = 3$

2 · Find the molality

m given

read it in mol/kg

from grams

$36.4 \text{ g} \rightarrow 0.699 \text{ m}$

3 · Multiply $i \cdot K \cdot m$

freezing: $K_f = 1.86$

$\Delta T_f = 3.90 \text{ }^\circ\text{C}$

boiling: $K_b = 0.512$

$\Delta T_b = i \cdot K_b \cdot m$

compare $i \times m$

same K : more particles, more shift

4 · Shift from pure water

freezing point

$0 - 3.90 = -3.90 \text{ }^\circ\text{C}$

boiling point

$100 \text{ }^\circ\text{C} + \Delta T_b$

the change asked

stop at step 3: report ΔT

🧠 **Sense check:** The route of worked example 3: all four steps, in order. ✓

Worked example 4: grams of salt for a boiling point

$$\Delta T_b = i \cdot K_b \cdot m, \text{ run backward}$$

given: 1.50 kg water · target boiling point 101.00 °C · NaCl 58.44 g/mol · $K_b = 0.512 \text{ °C} \cdot \text{kg/mol}$ ·
wanted: grams of NaCl

A cook wants a pot holding **1.50 kg** of water to boil at **101.00 °C**. Salt is the only solute. How many grams of NaCl must dissolve?

Run the route backward: undo the shift, undo the multiplication, then turn molality into grams.

✓ Worked example 4: solution

$$\Delta T_b = i \cdot K_b \cdot m, \text{ run backward}$$

$$1.50 \text{ kg water} \cdot \text{target } 101.00 \text{ }^\circ\text{C} \cdot \text{NaCl } 58.44 \text{ g/mol} \cdot K_b = 0.512 \text{ }^\circ\text{C}\cdot\text{kg/mol}$$

Step 1 · Count the particles

Undo the shift

$$\Delta T_b = 101.00 \text{ }^\circ\text{C} - 100.00 \text{ }^\circ\text{C} = 1.00 \text{ }^\circ\text{C}$$

Undo the multiplication

$$m = \frac{1.00 \text{ }^\circ\text{C}}{2 \times 0.512 \text{ }^\circ\text{C}\cdot\text{kg/mol}} = 0.9766 \text{ mol/kg}$$

🧠 **Sense check:** $i = 2$ and $m = 0.9766 \text{ mol/kg}$ carry into the grams. ✓

✓ Worked example 4: grams of salt

$m = 0.9766 \text{ mol NaCl per kg of water}$

$1.50 \text{ kg water} \cdot \text{NaCl } 58.44 \text{ g/mol} \cdot \text{wanted: grams of NaCl}$

Molality to grams

$$1.50 \text{ kg water} \times \frac{0.9766 \text{ mol NaCl}}{1 \text{ kg water}} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = 85.6 \text{ g NaCl}$$

 **Sense check:** Check forward: $85.6 \text{ g} \div 58.44 \text{ g/mol} \div 1.50 \text{ kg} = 0.977 \text{ m}$, and $2 \times 0.512 \times 0.9766 = 1.00 \text{ }^\circ\text{C}$. A pinch of salt in the pasta pot barely moves the boiling point. ✓

✓ Worked example 4: the route on the map

101.00 °C → ΔT_b 1.00 °C → m 0.9766 mol/kg → 1.465 mol → 85.6 g NaCl
the route run backward · i = 2 · Kb for boiling · kg of water, then molar mass

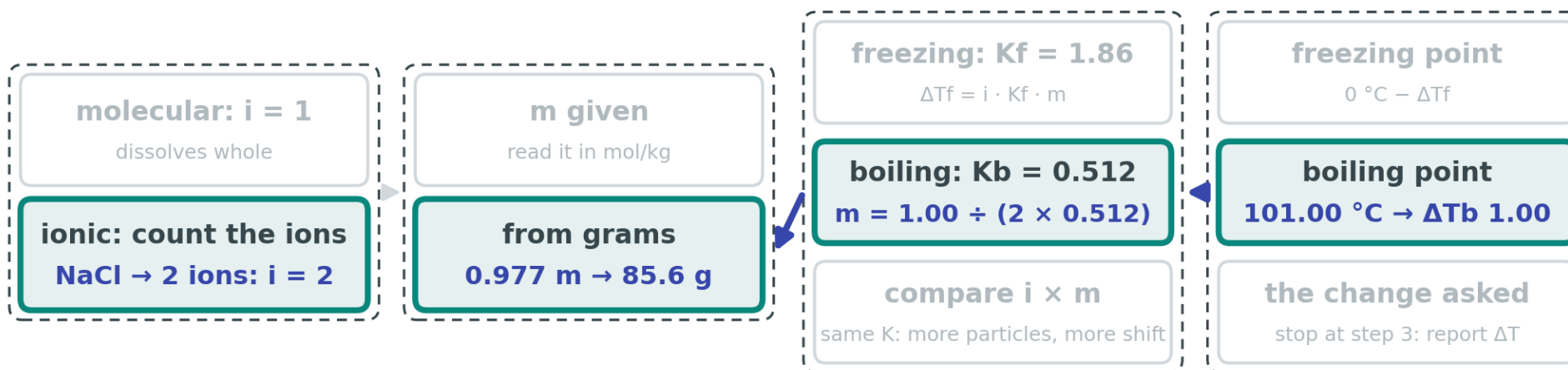
run backward: target temperature → ΔT → m → mol → grams · i from step 1 divides in step 3

1 · Count the particles

2 · Find the molality

3 · Multiply i · K · m

4 · Shift from pure water



Sense check: Each backward move undoes a forward one, in reverse order: subtract 100, divide by i · Kb, then use kg and molar mass. ✓



Practice 5

$$\Delta T_f = i \cdot K_f \cdot m$$

$$\text{molar mass MgCl}_2 \text{ 95.21 g/mol} \cdot K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$$

A road brine made with **600. g** of water must stay liquid down to **$-8.50 \text{ }^\circ\text{C}$** . How many grams of magnesium chloride, MgCl_2 , must dissolve?

- A. 0.914
- B. 261
- C. 87.0
- D. 37.5
- E. 316

✓ Practice 5 answer: C

$$\Delta T_f = i \cdot K_f \cdot m, \text{ run backward}$$

target $-8.50\text{ }^{\circ}\text{C}$: $\Delta T_f = 8.50\text{ }^{\circ}\text{C}$ · MgCl_2 gives $\text{Mg}^{2+} + 2\text{Cl}^-$: $i = 3$ · $600.\text{ g} = 0.600\text{ kg water}$

$$m = 8.50\text{ }^{\circ}\text{C} \div (3 \times 1.86\text{ }^{\circ}\text{C}\cdot\text{kg/mol}) = 1.523\text{ mol/kg} \rightarrow$$

$$0.600\text{ kg} \times 1.523\text{ mol/kg} = 0.914\text{ mol} \rightarrow 0.914\text{ mol} \times 95.21\text{ g/mol} = \mathbf{87.0\text{ g} \cdot \text{answer C}}$$

A stopped at moles: 0.914 mol is one step short of grams. B forgot i : $8.50 \div 1.86 \times 0.600 \times 95.21 = 261\text{ g}$. D flipped the rearrangement: $(3 \times 1.86) \div 8.50 \times 0.600 \times 95.21 = 37.5\text{ g}$. E used K_b : $8.50 \div (3 \times 0.512) \times 0.600 \times 95.21 = 316\text{ g}$.

🧠 **Sense check:** Check forward: $3 \times 1.86 \times 1.523 = 8.50\text{ }^{\circ}\text{C}$, so 87.0 g of MgCl_2 holds the freezing point at $-8.50\text{ }^{\circ}\text{C}$.



✓ Practice 5: the route on the map

$-8.50\text{ }^{\circ}\text{C} \rightarrow \Delta T_f\text{ }8.50\text{ }^{\circ}\text{C} \rightarrow m\text{ }1.523\text{ mol/kg} \rightarrow 0.914\text{ mol} \rightarrow 87.0\text{ g MgCl}_2$
the route run backward $\cdot i = 3 \cdot K_f$ for freezing $\cdot \text{kg of water, then molar mass}$

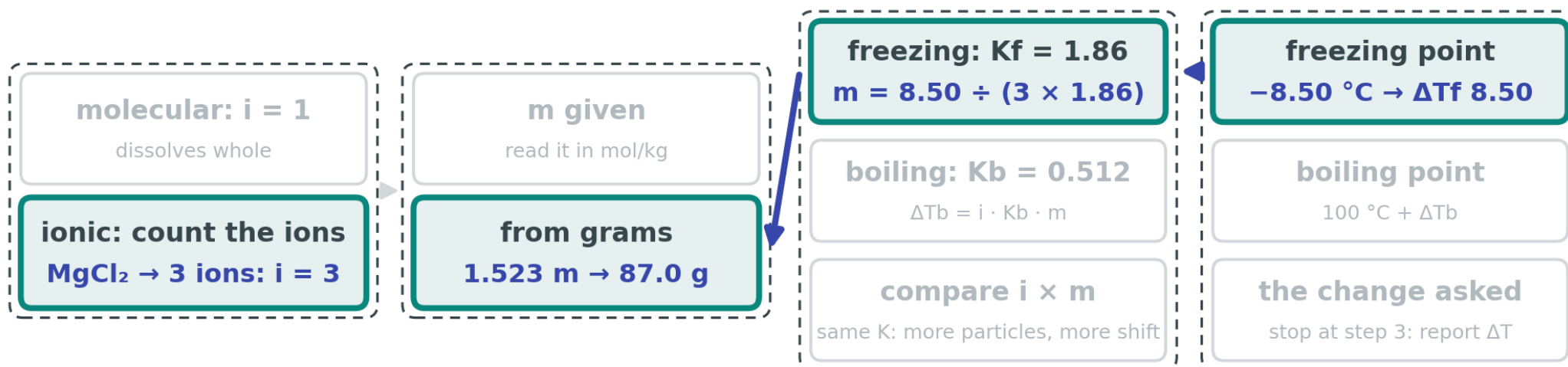
run backward: target temperature $\rightarrow \Delta T \rightarrow m \rightarrow \text{mol} \rightarrow \text{grams} \cdot i$ from step 1 divides in step 3

1 · Count the particles

2 · Find the molality

3 · Multiply $i \cdot K \cdot m$

4 · Shift from pure water



🧠 **Sense check:** The backward route of worked example 4, now with K_f and a freezing target. ✓



Practice 6

$$\Delta T_b = i \cdot K_b \cdot m$$

$K_b = 0.512\text{ }^{\circ}\text{C}\cdot\text{kg/mol}$ for every solution below

Which aqueous solution has the highest boiling point?

- A. sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$, at 0.75 m
- B. lithium bromide, LiBr , at 0.55 m
- C. sodium phosphate, Na_3PO_4 , at 0.25 m
- D. potassium nitrate, KNO_3 , at 0.45 m

✓ Practice 6 answer: B

$i \times m$: LiBr 1.10 > Na₃PO₄ 1.00 > KNO₃ 0.90 > sucrose 0.75

i: sucrose 1 · LiBr 2 (Li⁺ + Br⁻) · Na₃PO₄ 4 · KNO₃ 2 · LiBr 2 × 0.55 = 1.10 mol/kg of particles · answer B

A ranked by molality alone: sucrose stays whole, $i = 1$, so its 0.75 m is only 0.75 mol/kg of particles. C ranked by i alone: Na₃PO₄ gives 3 Na⁺ + PO₄³⁻, $i = 4$, but $4 \times 0.25 = 1.00$ falls short. D split the nitrate into atoms, $5 \times 0.45 = 2.25$; a polyatomic ion stays whole, so KNO₃ gives K⁺ and NO₃⁻, $2 \times 0.45 = 0.90$.

 **Sense check:** 1.10 mol/kg of particles is the most, so LiBr boils highest, near $100 + 2 \times 0.512 \times 0.55 = 100.56$ °C. Molality and i both count. ✓

 Practice 7

water: $K_f = 1.86\text{ }^{\circ}\text{C}\cdot\text{kg/mol}$ $K_b = 0.512\text{ }^{\circ}\text{C}\cdot\text{kg/mol}$

given: $1.70\text{ m Mg(NO}_3)_2$ · wanted: the boiling point

A **1.70 m** aqueous solution of magnesium nitrate, $\text{Mg(NO}_3)_2$, is heated at 1 atm. At what temperature, in $^{\circ}\text{C}$, does it boil?

- A. 102.61
- B. 100.87
- C. 97.39
- D. 90.51
- E. 2.61

✓ Practice 7 answer: A



given: $1.70 \text{ m} \cdot K_b = 0.512 \text{ }^\circ\text{C}\cdot\text{kg/mol}$ · each nitrate ion stays whole

$$\Delta T_b = 3 \times 0.512 \times 1.70 = 2.61 \text{ }^\circ\text{C} \rightarrow 100 \text{ }^\circ\text{C} + 2.61 \text{ }^\circ\text{C} = \mathbf{102.61 \text{ }^\circ\text{C} \cdot \text{answer A}}$$

B forgot i : $100 + 0.512 \times 1.70 = 100.87 \text{ }^\circ\text{C}$. C subtracted the rise: $100 - 2.61 = 97.39 \text{ }^\circ\text{C}$, but a solute raises the boiling point. D ran the freezing-point calculation: $100 - 3 \times 1.86 \times 1.70 = 90.51 \text{ }^\circ\text{C}$ uses K_f and a drop. E stopped at ΔT_b : $2.61 \text{ }^\circ\text{C}$ is the rise, not the boiling point.

 **Sense check:** Three ions per formula unit raise the boiling point $2.61 \text{ }^\circ\text{C}$ above pure water's $100 \text{ }^\circ\text{C}$. Boiling point goes UP. ✓



Check yourself

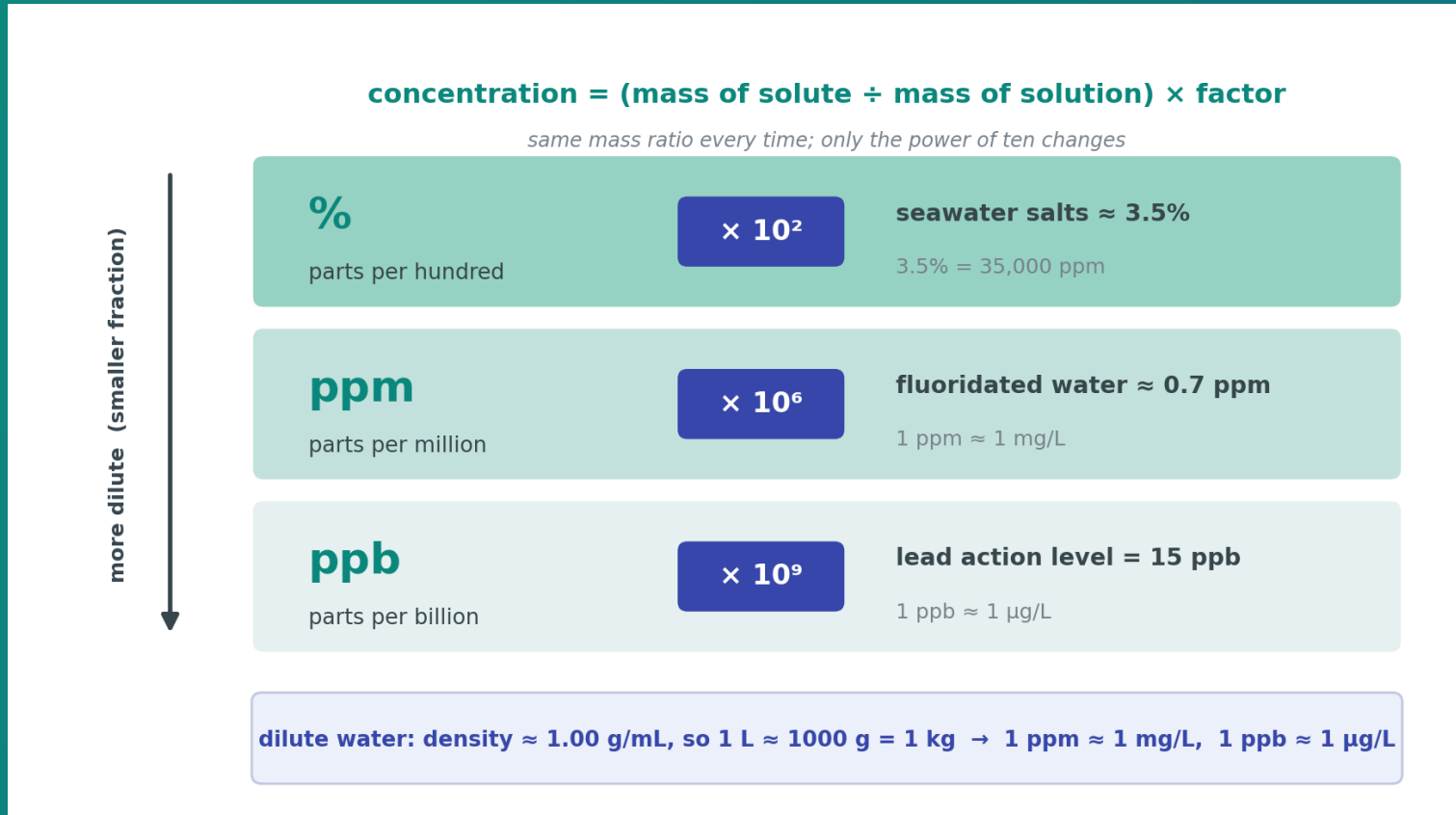
1. A 0.50 molal molecular-solute solution and a 0.50 molal NaCl solution are both cooled. Which freezes at the lower temperature, and why?
2. Write ΔT_b for a 0.750 molal Na₂SO₄ solution, then its boiling point. Which constant belongs in the formula?

Freezing and boiling points depend only on how many particles dissolve. The same particle count returns with acids and bases: a strong acid ionizes completely, a weak acid releases only a few ions.

9 · 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.

From mass percent to ppm

$$\text{mass percent} = (\text{g solute} \div \text{g solution}) \times 100$$

part over the whole solution, solute plus solvent

A **400. g** sample of spring water holds **0.0036 g of calcium**. Find the mass percent.

$$\frac{0.0036 \text{ g calcium}}{400. \text{ g solution}} \times 100 = 0.00090\%$$

A percent this small is hard to read. The same fraction $\times 10^6$ reads **9.0 ppm**.

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. 1 ppm = 1 mg/L is given with every problem that uses it

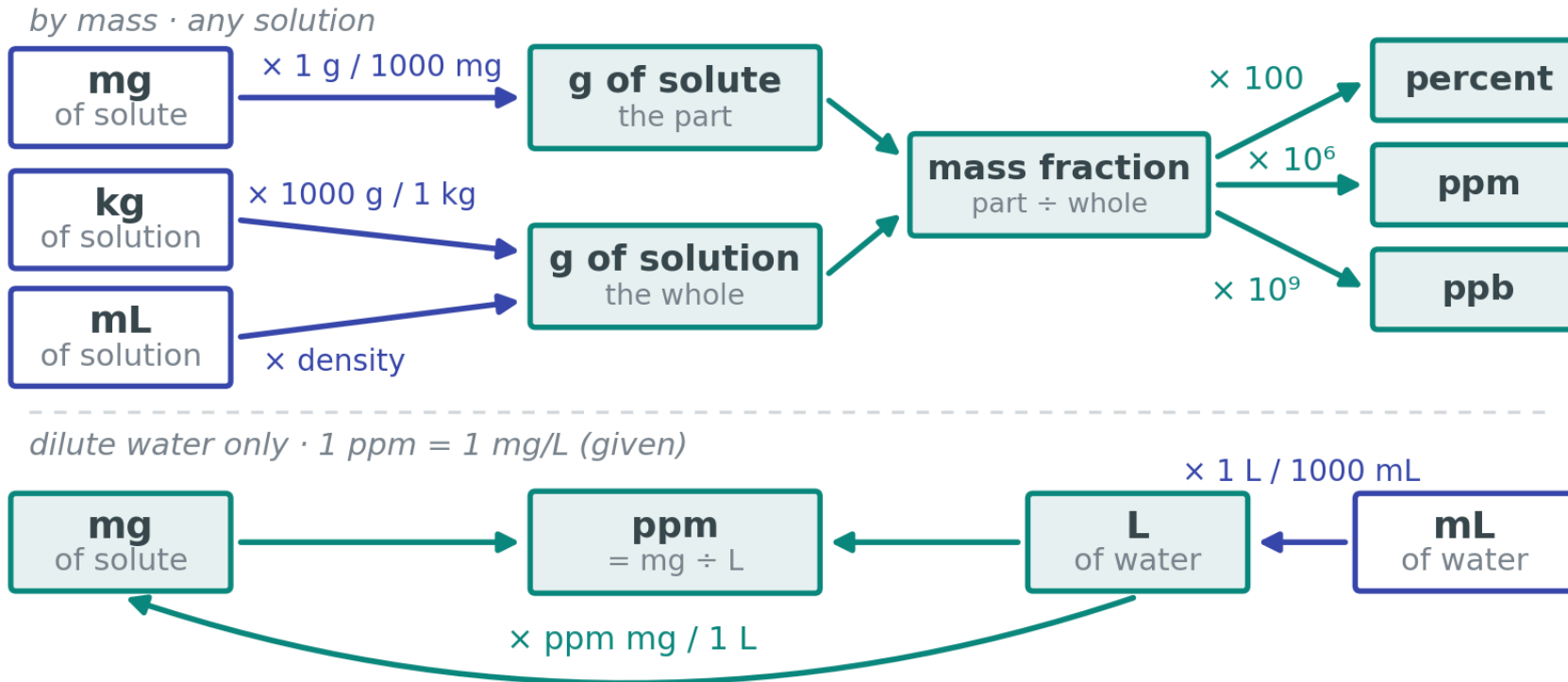
A ppm then reads as mg/L. A dense brine or a nonaqueous solvent breaks the shortcut.



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.

One route for ppm and ppb



Any solution: both masses in grams, part over whole, then one power of ten. Dilute water only: with the given 1 ppm = 1 mg/L, milligrams over liters is already ppm.



Guided example: chromium in river water

37.0 mg chromium in a 375.0 g sample of river water

given: 37.0 mg solute · 375.0 g solution · wanted: mass percent, ppm, ppb

A **375.0 g** sample of river water taken near an industrial plant contains **37.0 mg of chromium**. What is the chromium concentration as a mass percent, in ppm, and in ppb?

The solute is in milligrams and the solution in grams. Match the units first.

✓ Guided example: solution

37.0 mg chromium in a 375.0 g sample of river water

given: 37.0 mg solute · 375.0 g solution · wanted: mass percent, ppm, ppb

Step 1 · Mass fraction, same units

$$37.0 \text{ mg Cr} \times \frac{1 \text{ g}}{1000 \text{ mg}} = 0.0370 \text{ g Cr} \quad \frac{0.0370 \text{ g Cr}}{375.0 \text{ g solution}} = 9.87 \times 10^{-5}$$

Step 2 · Multiply by the power of ten

$$\times 100 = 9.87 \times 10^{-3} \% \cdot \times 10^6 = 98.7 \text{ ppm} \cdot \times 10^9 = 9.87 \times 10^4 \text{ ppb}$$

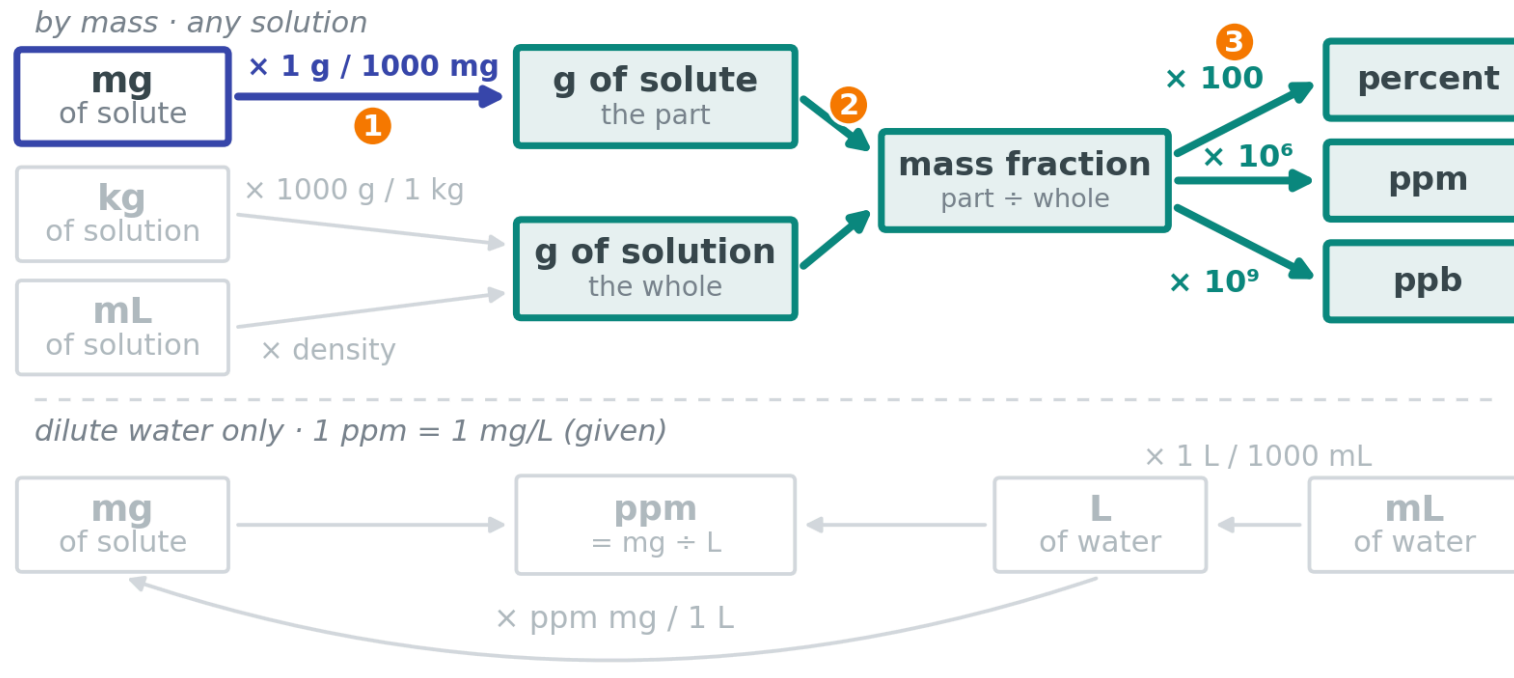
each power of ten multiplies the same fraction, 9.87×10^{-5} : the same chromium on three scales

🧠 **Sense check:** 98.7 ppm is 9.87×10^4 ppb: each step down the ladder is $\times 1000$. The percent, 0.00987%, is the hardest of the three to read. ✓

✓ Guided example: the route on the map

37.0 mg chromium in a 375.0 g sample of river water

given: 37.0 mg solute · 375.0 g solution · found: $9.87 \times 10^{-3} \%$ · 98.7 ppm · 9.87×10^4 ppb



🧠 **Sense check:** Three moves: milligrams to grams, part over whole, then one power of ten for each unit. ✓



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

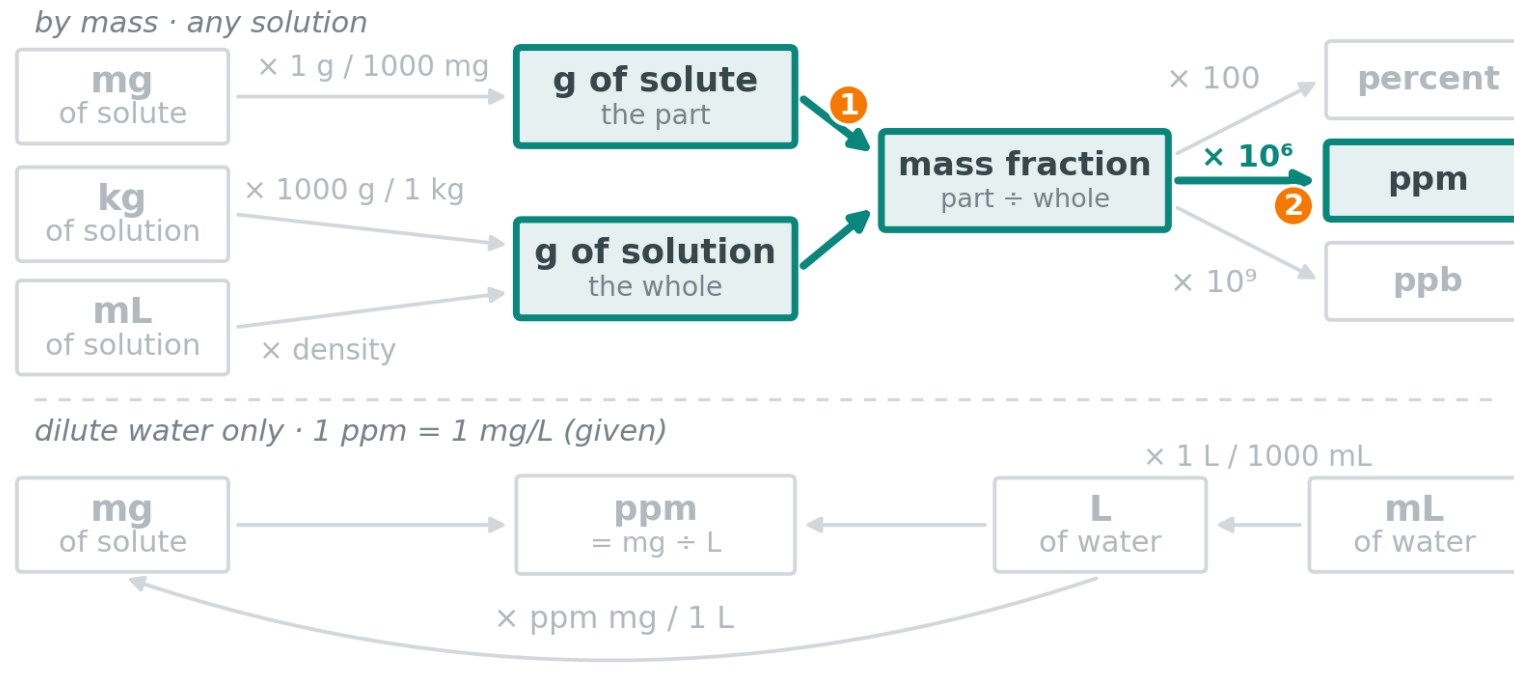
$$1.2 \times 10^{-5} \times 10^6 = 12 \text{ ppm}$$
$$0.0030 / 250 \times 10^6 = 12 \text{ ppm of copper}$$

🧠 **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. ✓

 Practice 1

swordfish fillet: 1.8×10^{-4} g mercury in 200.0 g
given: 1.8×10^{-4} g solute · 200.0 g sample · wanted: ppm

A **200.0 g** swordfish fillet contains **1.8×10^{-4} g** of mercury. What is the mercury concentration, in ppm?

- A. 900
- B. 9.0×10^{-7}
- C. 0.90
- D. 9.0×10^{-5}

✓ Practice 1: answer C

$$\text{ppm} = (\text{mass solute} / \text{mass solution}) \times 10^6$$

$1.8 \times 10^{-4} \text{ g mercury} \cdot 200.0 \text{ g fillet} \cdot \text{both already in grams}$

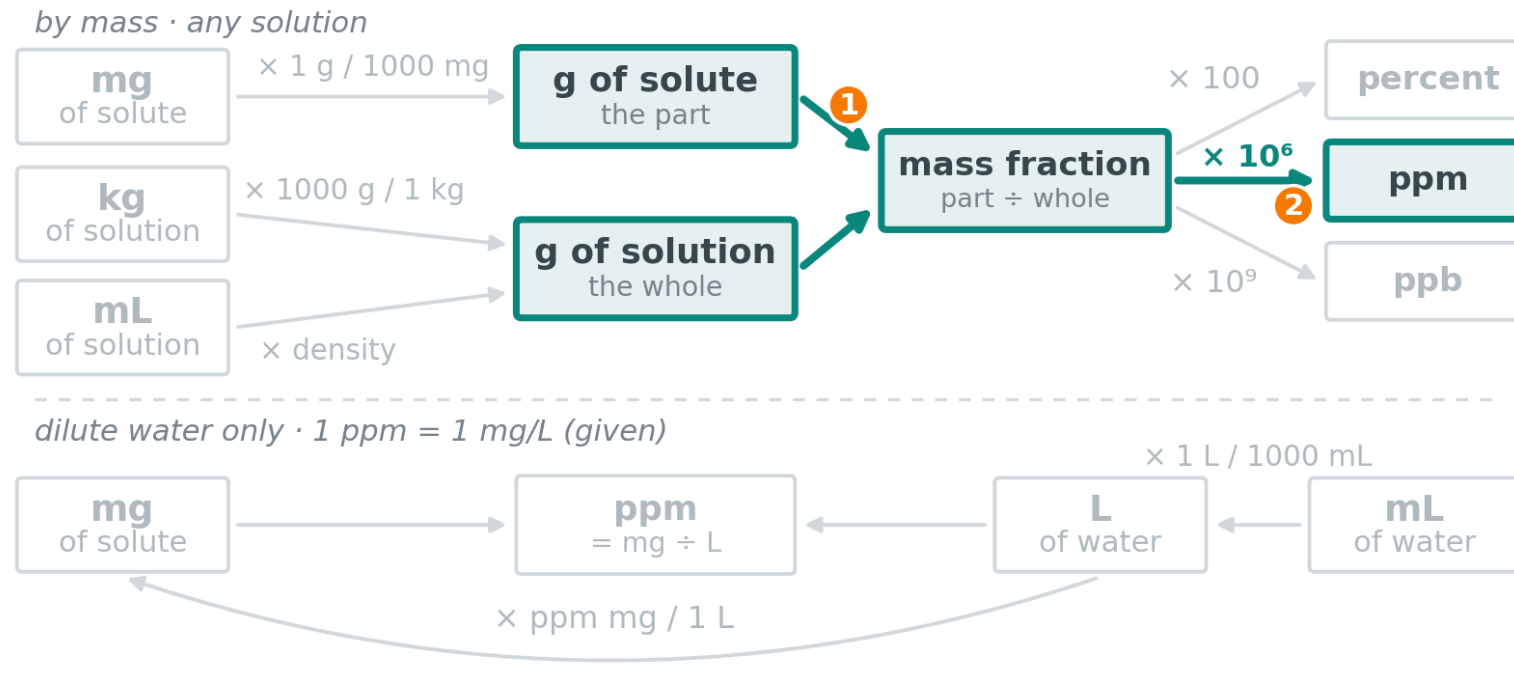
$$\frac{1.8 \times 10^{-4} \text{ g Hg}}{200.0 \text{ g fillet}} \times 10^6 = \mathbf{0.90 \text{ ppm (answer C)}}$$

A used 10^9 : $1.8 \times 10^{-4} \div 200.0 \times 10^9 = 900$ is ppb. B stopped at the mass fraction: $1.8 \times 10^{-4} \div 200.0 = 9.0 \times 10^{-7}$. D used $\times 100$: 9.0×10^{-5} is the mass percent.

 **Sense check:** 0.90 mg of mercury in each kilogram of fish, just under the FDA's 1 ppm action level for mercury in fish. ✓

✓ Practice 1: the route on the map

swordfish fillet: 1.8×10^{-4} g mercury in 200.0 g
 given: 1.8×10^{-4} g solute · 200.0 g sample · found: 0.90 ppm



🧠 **Sense check:** Both masses were already in grams: part over whole, then $\times 10^6$. ✓

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.

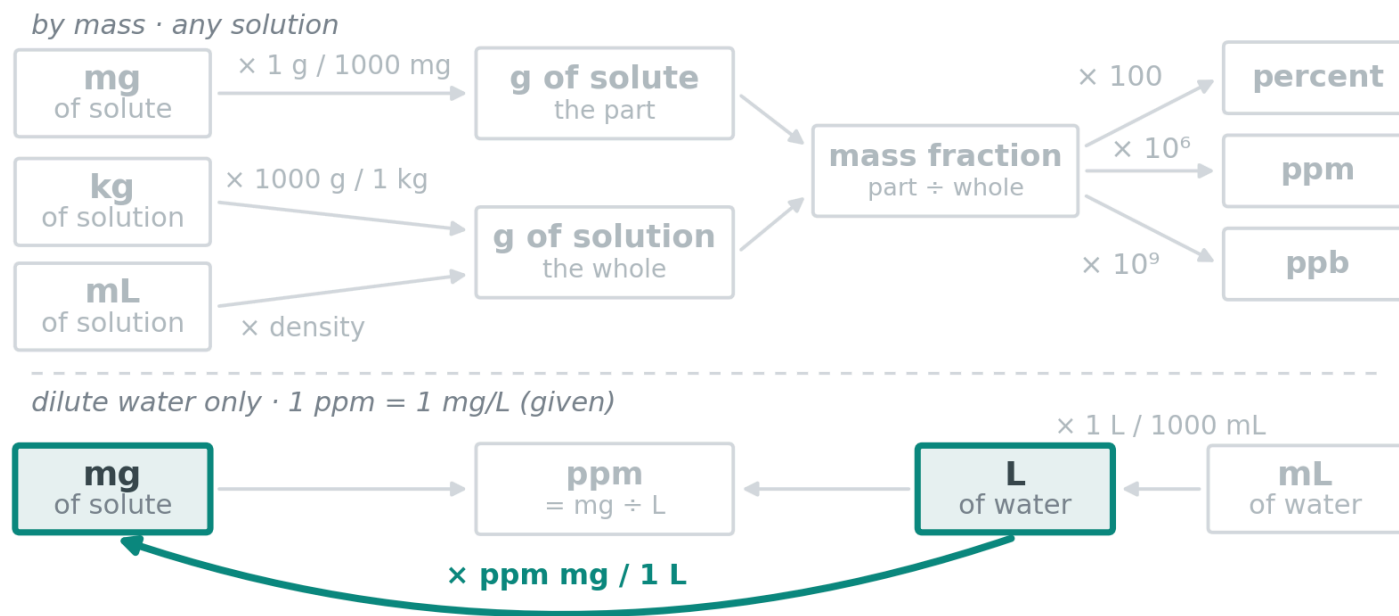
$$2.0 \text{ L} \times \frac{\text{mg F}^-}{1 \text{ L}} = \text{mg F}^-$$

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

✓ Your turn: the route on the map

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

given: 0.70 ppm $\cdot$ 2.0 L of water $\cdot$ found: 1.4 mg F⁻



 **Sense check:** Dilute water, so the ppm reads as milligrams per liter. One factor carries liters of water to milligrams of fluoride. ✓

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 2

well water: 0.84 mg nitrate in 2.40 L
given: 1 ppm = 1 mg/L for dilute water solutions

A **2.40 L** jug of well water contains **0.84 mg** of dissolved nitrate. What is the nitrate concentration, in ppm?

- A. 0.35
- B. 350
- C. 2.9
- D. 2.0

✓ Practice 2: answer A

1 ppm = 1 mg/L for dilute water solutions (given)

0.84 mg nitrate · 2.40 L of well water · wanted: ppm

$$\frac{0.84 \text{ mg nitrate}}{2.40 \text{ L water}} = 0.35 \text{ mg/L} = \mathbf{0.35 \text{ ppm (answer A)}}$$

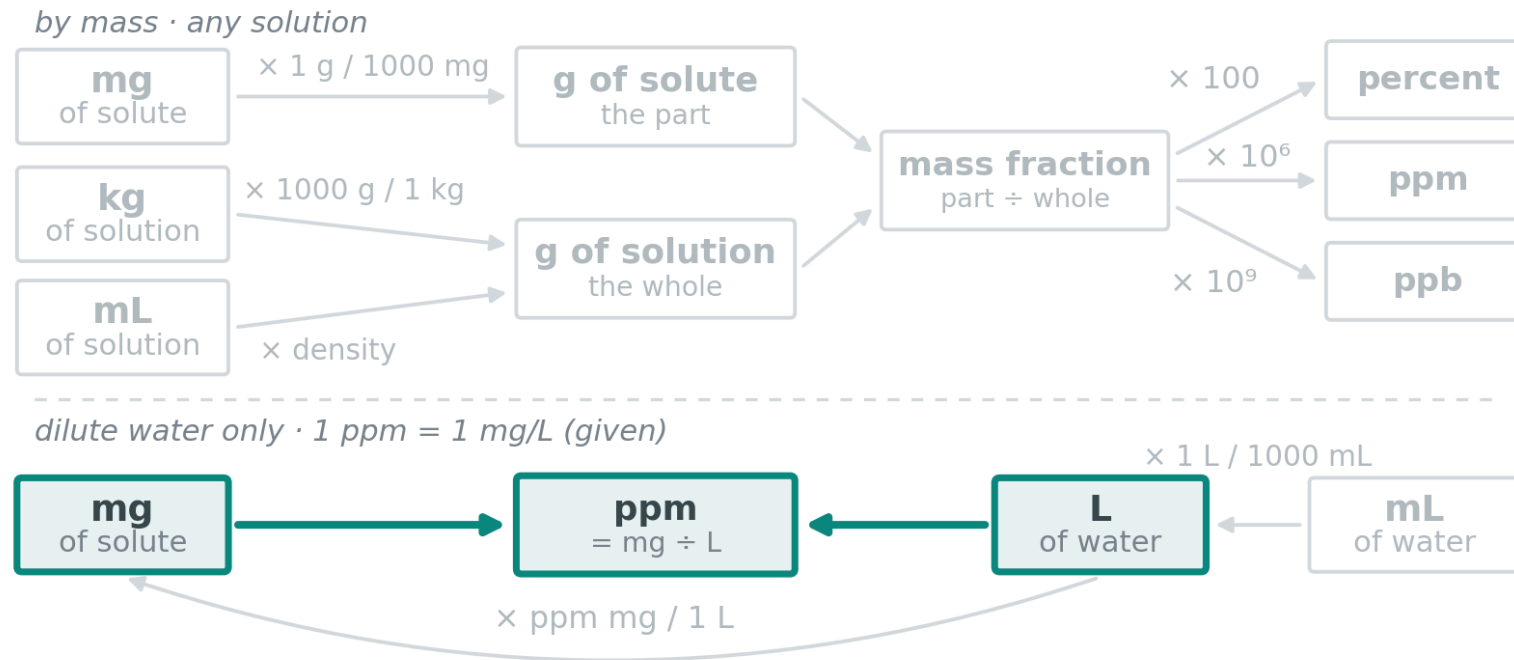
B kept milligrams over grams: 2.40 L is about 2400 g, and $0.84 \div 2400 \times 10^6 = 350$, a thousand times too high. C flipped the fraction: $2.40 \div 0.84 = 2.9$. D multiplied: $0.84 \times 2.40 = 2.0$.

🧠 **Sense check:** Less than 1 mg spread through more than 2 L must land below 1 mg/L: 0.35 ppm. ✓

✓ Practice 2: the route on the map

well water: 0.84 mg nitrate in 2.40 L

given: 1 ppm = 1 mg/L · found: 0.35 ppm



 **Sense check:** Dilute water with milligrams and liters in hand: one move, milligrams over liters. ✓



Practice 3

groundwater: 0.036 mg manganese in a 150. mL sample

given: 1 ppm = 1 mg/L for dilute water solutions

A **150. mL** sample of groundwater contains **0.036 mg** of manganese. What is the manganese concentration, in ppm?

- A. 240
- B. 0.0054
- C. 4.2
- D. 2.4×10^{-4}
- E. 0.24

✓ Practice 3: answer E

1 ppm = 1 mg/L for dilute water solutions (given)
 0.036 mg manganese · 150. mL of groundwater · wanted: ppm

Two conversion factors are needed.

$$150. \cancel{\text{mL}} \times \frac{1 \text{ L}}{1000 \cancel{\text{mL}}} = 0.150 \text{ L} \quad \frac{0.036 \text{ mg Mn}}{0.150 \text{ L water}} = \mathbf{0.24 \text{ ppm (answer E)}}$$

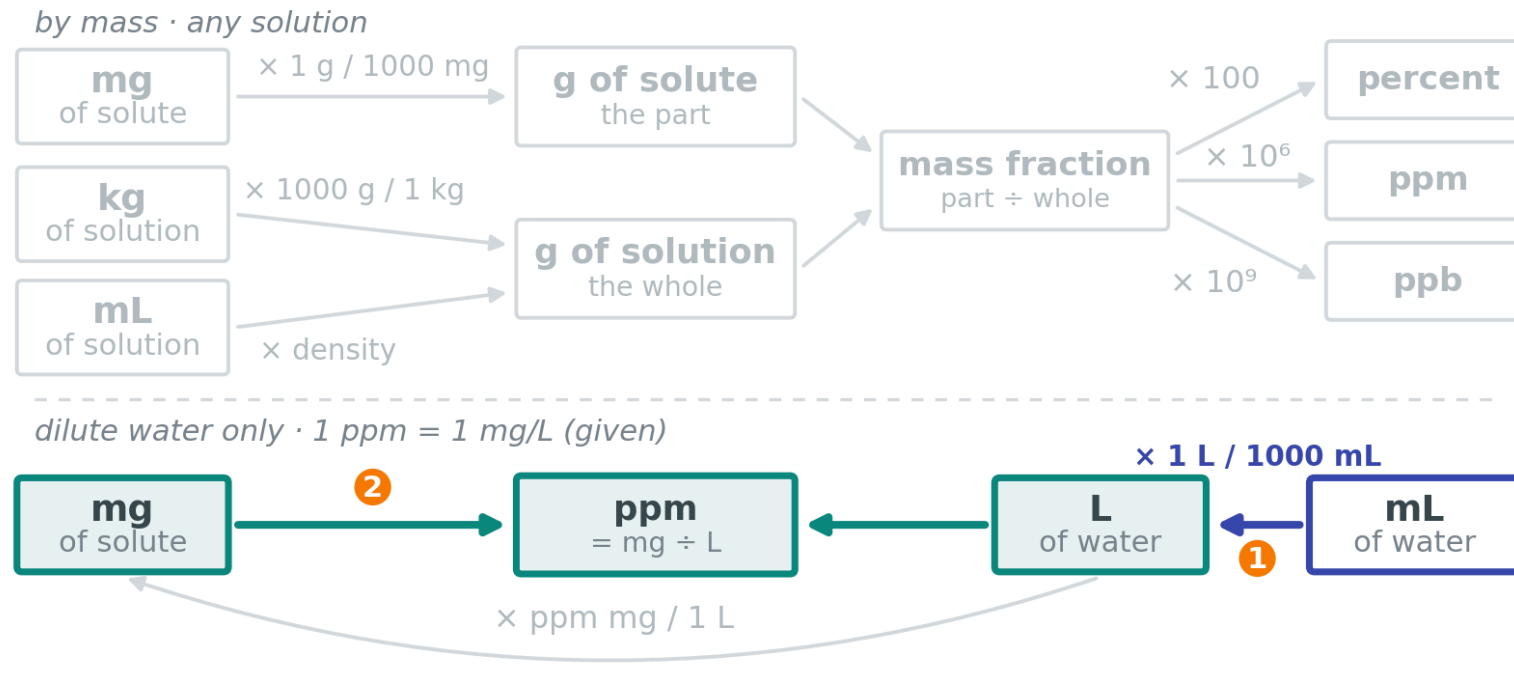
A kept milligrams over grams: 150. mL is about 150. g, and $0.036 \div 150. \times 10^6 = 240$. B multiplied: $0.036 \times 0.150 = 0.0054$. C flipped: $0.150 \div 0.036 = 4.2$. D skipped milliliters to liters: $0.036 \div 150. = 2.4 \times 10^{-4}$.

 **Sense check:** A liter is about 6.67 times 150. mL, so a full liter of this water holds $0.036 \times 6.67 = 0.24 \text{ mg}$. ✓

✓ Practice 3: the route on the map

groundwater: 0.036 mg manganese in a 150. mL sample

given: 1 ppm = 1 mg/L · found: 0.150 L · 0.24 ppm



🧠 **Sense check:** Two moves: milliliters to liters first, then milligrams over liters. ✓



Practice 4

lake water: 0.024 mg lead in a 1.50 kg sample
given: 0.024 mg solute · 1.50 kg solution · wanted: ppb

A **1.50 kg** sample of lake water contains **0.024 mg** of lead. What is the lead concentration, in ppb?

- A. 0.016
- B. 1.6×10^{-8}
- C. 1.6×10^{-6}
- D. 16

✓ Practice 4: answer D

$$\text{ppb} = (\text{mass solute} / \text{mass solution}) \times 10^9$$

0.024 mg = 2.4×10^{-5} g lead · 1.50 kg = 1500 g solution · both in grams first

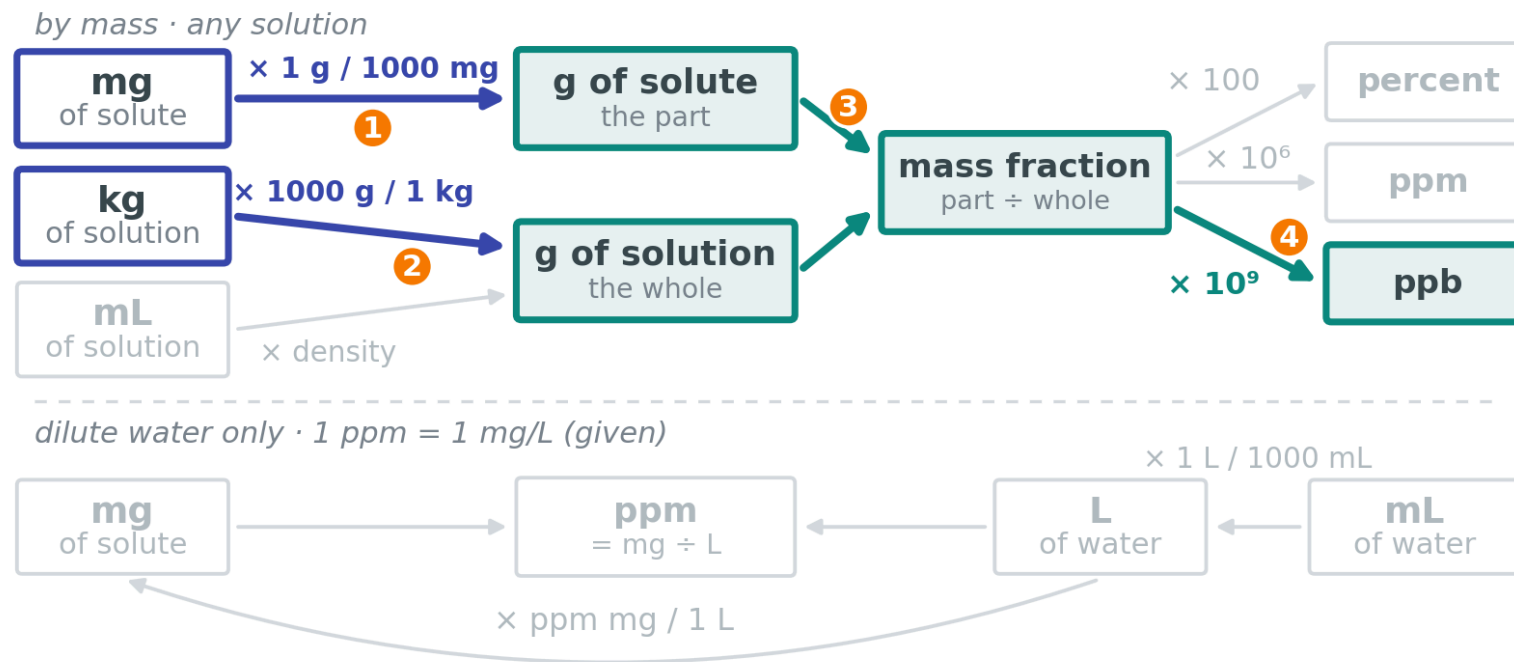
$$\frac{2.4 \times 10^{-5} \text{ g Pb}}{1500 \text{ g solution}} \times 10^9 = \mathbf{16 \text{ ppb (answer D)}}$$

A used 10^6 : 0.016 is the same lead in ppm. B stopped at the mass fraction: $2.4 \times 10^{-5} \div 1500 = 1.6 \times 10^{-8}$. C used $\times 100$: 1.6×10^{-6} is the mass percent.

🧠 **Sense check:** 16 ppb sits just above the 15 ppb action level set for lead in drinking water. ✓

✓ Practice 4: the route on the map

lake water: 0.024 mg lead in a 1.50 kg sample
given: 0.024 mg solute · 1.50 kg solution · found: 16 ppb



🧠 **Sense check:** Four moves: milligrams to grams, kilograms to grams, part over whole, then $\times 10^9$ for ppb. ✓

 Practice 5

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 5: 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

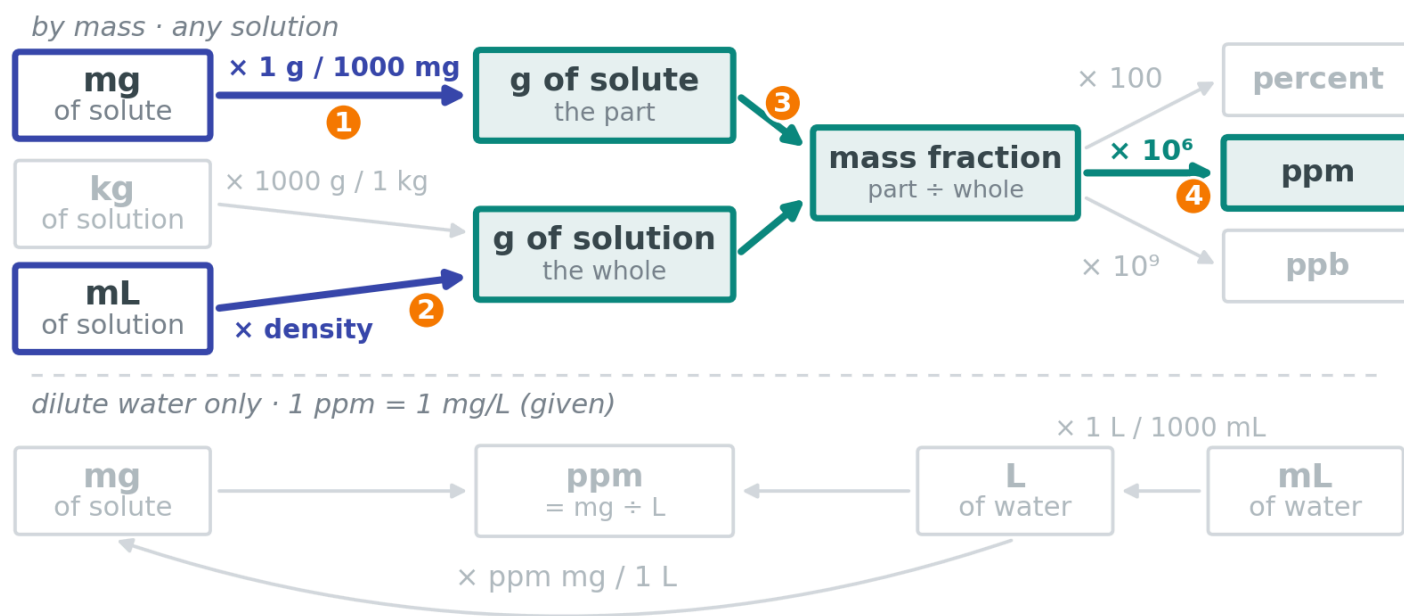
$$\frac{0.00416 \text{ g tin}}{208.0 \text{ g solution}} \times 10^6 = \mathbf{20.0 \text{ ppm (answer B)}}$$

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. ✓

✓ Practice 5: the route on the map

4.16 mg tin in 160.0 mL of concentrate · density 1.30 g/mL
 given: 4.16 mg solute · 160.0 mL solution · found: 0.00416 g · 208.0 g · 20.0 ppm



Sense check: Not dilute water, so the mass route: milligrams to grams, milliliters to grams through the density, part over whole, then $\times 10^6$. ✓

Extra practice 1

garden pond: 0.30 ppm copper · 54 mg copper per packet
given: 1 ppm = 1 mg/L for dilute water solutions

A garden-pond algae treatment is dosed to **0.30 ppm** of copper. One packet holds **54 mg** of copper. How many liters of pond water does one packet treat?

- A. 16
- B. 0.0056
- C. 0.18
- D. 1.8×10^2

✓ Extra practice 1: answer D

1 ppm = 1 mg/L for dilute water solutions (given)

0.30 ppm = 0.30 mg copper in each liter · 54 mg copper per packet · wanted: liters of water

$$54 \text{ mg Cu} \times \frac{1 \text{ L water}}{0.30 \text{ mg Cu}} = 1.8 \times 10^2 \text{ L (answer D)}$$

A multiplied: $54 \times 0.30 = 16$, and $\text{mg} \times \text{mg/L}$ leaves mg^2/L , not liters. B flipped the factor: $0.30 \div 54 = 0.0056$. C changed the dose to grams but kept 0.30 mg/L: $0.054 \div 0.30 = 0.18$, and grams over milligrams do not cancel.

🧠 **Sense check:** Each liter holds only 0.30 mg, so 54 mg spreads through far more than 54 L of water: $1.8 \times 10^2 \text{ L}$. ✓



Extra practice 2

well water: 6.0 ppb arsenic · 2.5 L a day · 60 days

given: 1 ppb = 1 µg/L for dilute water solutions

A well-water report lists arsenic at **6.0 ppb**. A resident drinks **2.5 L** of this water each day. How many milligrams of arsenic does the resident take in over **60 days**?

- A. 900
- B. 0.015
- C. 9.0×10^{-4}
- D. 0.36
- E. 0.90

✓ Extra practice 2: answer E

1 ppb = 1 µg/L for dilute water solutions (given)

6.0 ppb = 6.0 µg arsenic in each liter · 2.5 L a day · 60 days · wanted: mg of arsenic

$$60 \text{ d} \times \frac{2.5 \text{ L}}{1 \text{ d}} \times \frac{6.0 \text{ µg As}}{1 \text{ L}} \times \frac{1 \text{ mg}}{1000 \text{ µg}} = \mathbf{0.90 \text{ mg As (answer E)}}$$

A stopped in micrograms: $60 \times 2.5 \times 6.0 = 900 \text{ µg}$, labeled mg. B left out the days: $2.5 \times 6.0 \div 1000 = 0.015 \text{ mg}$ is one day. C divided by 10^6 : $900 \text{ µg} \div 10^6 = 9.0 \times 10^{-4}$ is grams, not milligrams. D left out the volume: $60 \times 6.0 \div 1000 = 0.36$ reads 6.0 ppb as 6.0 µg a day.

 **Sense check:** One day's water carries $2.5 \times 6.0 = 15 \text{ µg}$. Sixty days carry 900 µg, just under one milligram. ✓



Extra practice 3

Na_2SiF_6 : sodium fluorosilicate

given: 8.00 g Na_2SiF_6 · 6.00×10^3 L of water · 1 ppm = 1 mg/L for dilute water solutions

A small water system fluoridates a 6.00×10^3 L storage tank with 8.00 g of sodium fluorosilicate, Na_2SiF_6 . What is the fluoride concentration, in ppm?

- A. 0.808
- B. 1.33
- C. 4.85×10^3
- D. 808
- E. 2.20

✓ Extra practice 3: answer A

$$\text{Na}_2\text{SiF}_6: 2(22.99) + 28.09 + 6(19.00) = 188.07 \text{ g/mol}$$

$$8.00 \text{ g Na}_2\text{SiF}_6 \cdot 6.00 \times 10^3 \text{ L of water} \cdot 1 \text{ ppm} = 1 \text{ mg/L (given)} \cdot \text{wanted: ppm F}$$

$$8.00 \text{ g Na}_2\text{SiF}_6 \times \frac{114.00 \text{ g F}}{188.07 \text{ g Na}_2\text{SiF}_6} \times \frac{1000 \text{ mg}}{1 \text{ g}} = 4.85 \times 10^3 \text{ mg F}$$

$$4.85 \times 10^3 \text{ mg F} \div 6.00 \times 10^3 \text{ L water} = 0.808 \text{ mg/L} = \mathbf{0.808 \text{ ppm (answer A)}}$$

B used the whole compound: $8.00 \times 10^3 \text{ mg} \div 6.00 \times 10^3 \text{ L} = 1.33$. C stopped at $4.85 \times 10^3 \text{ mg}$ of fluoride. D took liters as grams: $4.85 \text{ g} \div 6.00 \times 10^3 \text{ g} \times 10^6 = 808$. E flipped the mass fraction: $8.00 \times 188.07 \div 114.00 = 13.2 \text{ g}$, so $13.2 \times 10^3 \text{ mg} \div 6.00 \times 10^3 \text{ L} = 2.20$.

🧠 **Sense check:** Fluoride is 60.6% of the compound, so it sits below 1.33 mg/L: 0.808 ppm. ✓

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, a solution's volume and molarity carry through a balanced equation: solution stoichiometry.

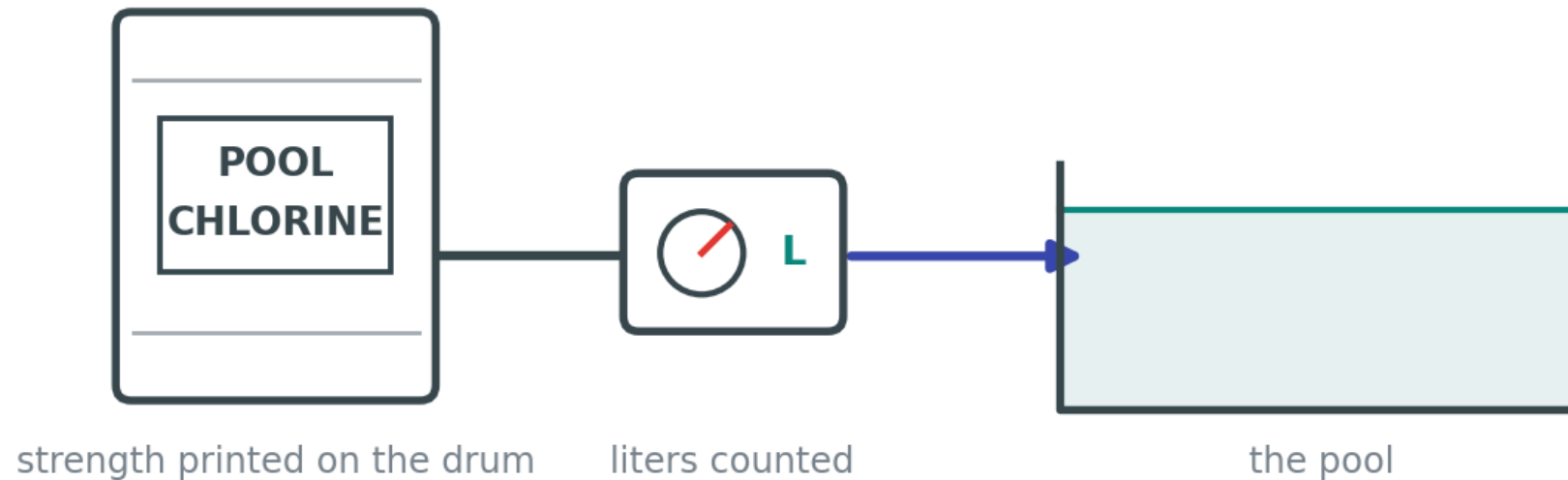
10 · Solution Stoichiometry

Convert a volume of solution to moles with molarity, cross substances with the mole ratio, and finish in grams or in the volume of a second solution.

Dosed by volume

dose = liters pumped × strength on the label

nothing is weighed



A pool is chlorinated by pumping in a measured volume of chlorine solution. Nothing is weighed: the strength on the drum's label and the liters pumped set the chlorine dose.

Molarity as a conversion factor

1.50 M HCl: 1.50 mol HCl = 1 L of solution

liters → moles: mol over L · moles → liters: L over mol

A solution's volume and molarity give the moles of a reactant or product. Either orientation works:

$$\frac{1.50 \text{ mol HCl}}{1 \text{ L soln}} \text{ or } \frac{1 \text{ L soln}}{1.50 \text{ mol HCl}}$$

What volume carries 3.00 mol HCl?

$$3.00 \text{ mol HCl} \times \frac{1 \text{ L soln}}{1.50 \text{ mol HCl}} = \mathbf{2.00 \text{ L soln}}$$

The mole ratio then crosses to another substance.

A solution's label counts moles

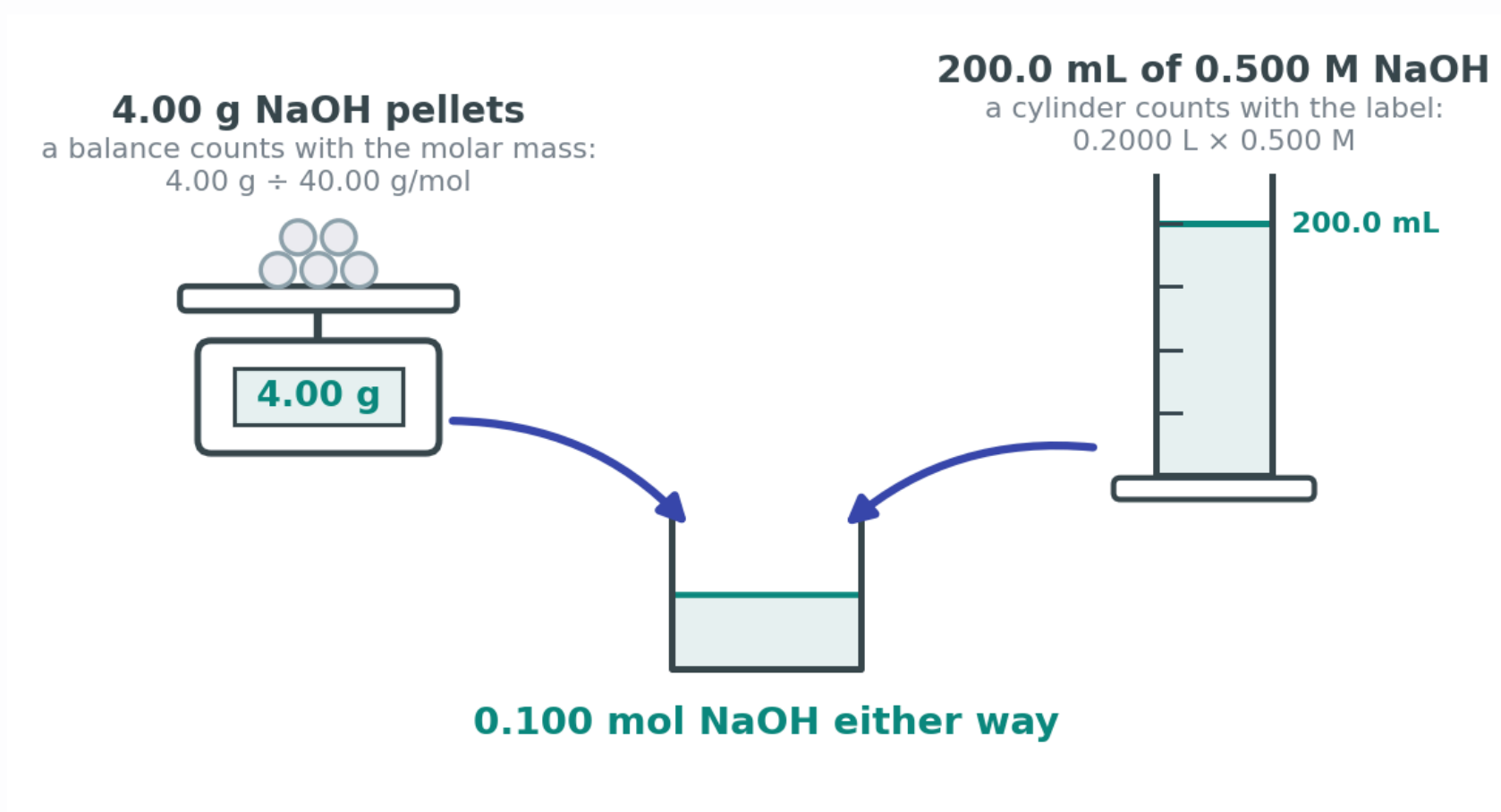
0.500 M NaOH: 0.500 mol NaOH = 1 L of solution
the label states a rate: moles delivered per liter poured

Measure a volume, and the label converts it to moles; the liters cancel:

$$0.2000 \text{ L soln} \times \frac{0.500 \text{ mol NaOH}}{1 \text{ L soln}} = \mathbf{0.100 \text{ mol NaOH}}$$

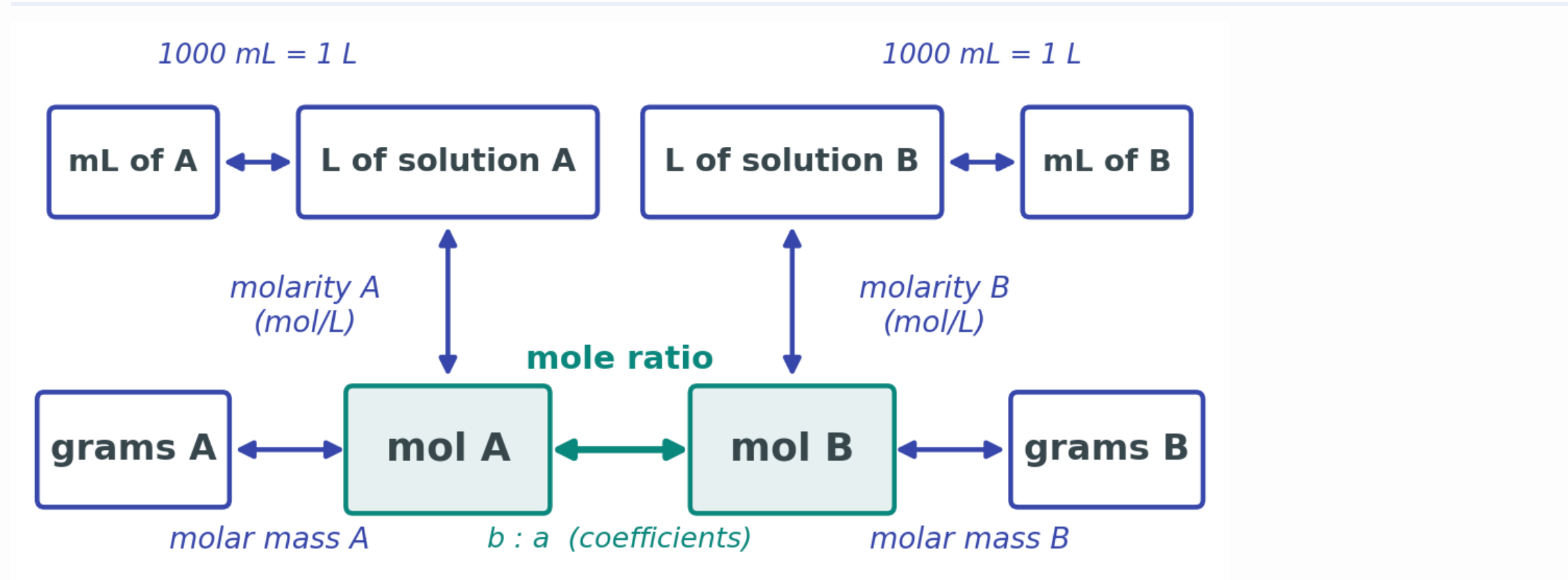
A graduated cylinder now counts moles. No balance is needed.

Two ways to deliver 0.100 mol



A reaction receives 0.100 mol either way. Stoichiometry works on moles; where they came from never enters the calculation.

Volume joins the map



Volume of solution enters through molarity, exactly where grams enter through molar mass. Every route still crosses the mole bridge, and the mole ratio still switches substances.



The method

1. **Volume → moles:** mL to L, then molarity. For grams, the molar mass.
2. **Moles → moles:** cross substances with the mole ratio. No other step can.
3. **Moles → the wanted unit:** molar mass for grams; molarity for volume.



Guided example: silver chloride



given: 0.500 L of 0.100 M NaCl · wanted: mol AgCl

Silver nitrate solution is added to **0.500 L of 0.100 M NaCl** until no more white AgCl forms. How many moles of AgCl precipitate?

Check each method step in turn. The volume is already in liters.

✓ Guided example: solution



given: 0.500 L of 0.100 M NaCl · wanted: mol AgCl

Step 1 · Volume → moles

Step 2 · Moles → moles

$$0.500 \text{ L soln} \times \frac{0.100 \text{ mol NaCl}}{1 \text{ L soln}} \times \frac{1 \text{ mol AgCl}}{1 \text{ mol NaCl}} = 0.0500 \text{ mol AgCl}$$

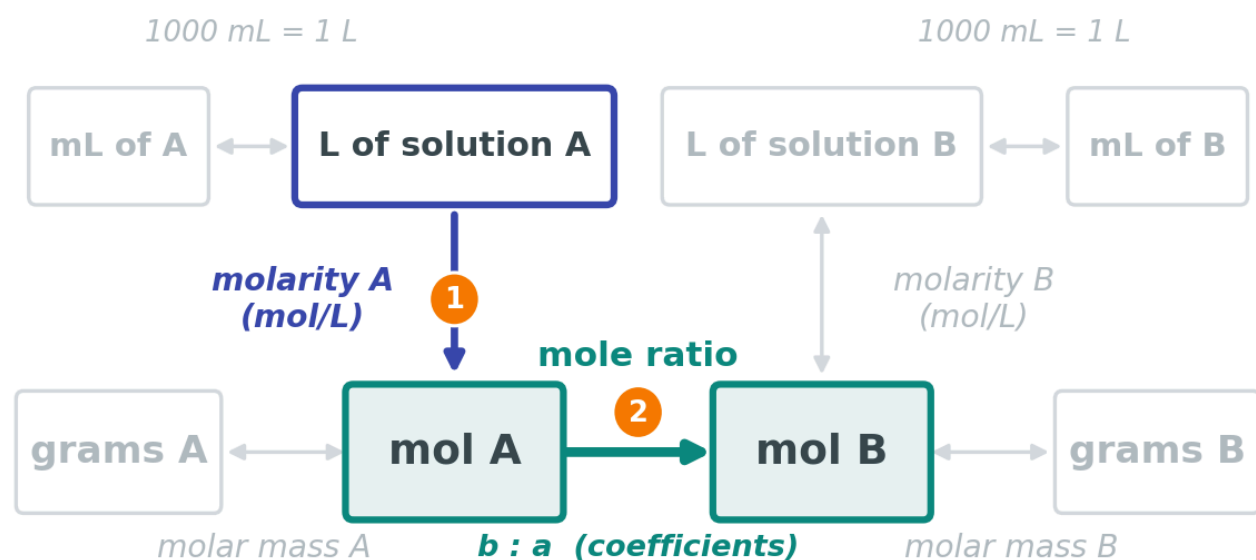
Step 3 · Moles → the wanted unit

 **Sense check:** One NaCl gives one AgCl, so the moles pass through the ratio unchanged: 0.0500 mol NaCl → 0.0500 mol AgCl ✓

✓ Guided example: the route on the map



given: 0.500 L of 0.100 M NaCl · found: 0.0500 mol AgCl



🧠 **Sense check:** A is NaCl, B is AgCl. The volume was already in liters, so the route starts at L of solution A: arrow 1 is the molarity, arrow 2 the mole ratio. ✓

 Practice 1

Hard-water scale is calcium carbonate. A lab mixes **40.0 mL of 0.250 M Na₂CO₃** with excess CaCl₂. How many moles of CaCO₃ precipitate?

- A. 0.0200
- B. 0.0100
- C. 0.160
- D. 10.0

✓ Practice 1: answer B



given: 40.0 mL of 0.250 M Na_2CO_3 · wanted: mol CaCO_3

$$0.0400 \text{ L soln} \times \frac{0.250 \text{ mol Na}_2\text{CO}_3}{1 \text{ L soln}} \times \frac{1 \text{ mol CaCO}_3}{1 \text{ mol Na}_2\text{CO}_3} = 0.0100 \text{ mol CaCO}_3 \text{ (answer B)}$$

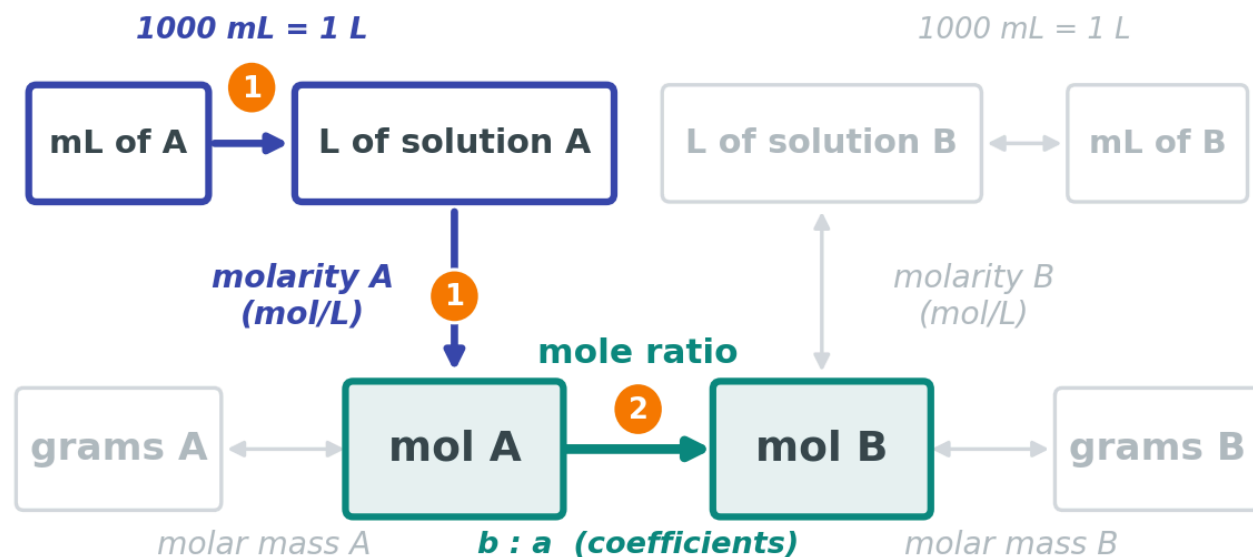
A read NaCl's coefficient for CaCO_3 : $0.0100 \times 2 = 0.0200$. C inverted the molarity: $0.0400 \div 0.250 = 0.160$. D fed milliliters to the molarity: $40.0 \times 0.250 = 10.0$.

🧠 **Sense check:** One CaCO_3 forms per Na_2CO_3 , so the 1 : 1 ratio leaves the moles unchanged: 0.0100 mol ✓

✓ Practice 1: the route on the map



given: 40.0 mL of 0.250 M Na_2CO_3 · found: 0.0100 mol CaCO_3



🧠 **Sense check:** A is Na_2CO_3 , B is CaCO_3 . Step 1 is two arrows: mL → L, then the molarity. D skipped the first one and fed milliliters to the molarity. ✓



Worked example 1: moles from a solution volume



given: 250.0 mL of 0.400 M HCl · wanted: mol H₂

Zinc metal dissolves in hydrochloric acid, releasing hydrogen gas. **250.0 mL of 0.400 M HCl** reacts completely with excess zinc. How many moles of H₂ form?

Write the route first: mL → L → mol HCl → mol H₂.

✓ Worked example 1: solution



given: 250.0 mL of 0.400 M HCl · wanted: mol H₂

Step 1 · Volume → moles

$$0.2500 \text{ L soln} \times \frac{0.400 \text{ mol HCl}}{1 \text{ L soln}} = 0.100 \text{ mol HCl}$$

Step 2 · Moles → moles

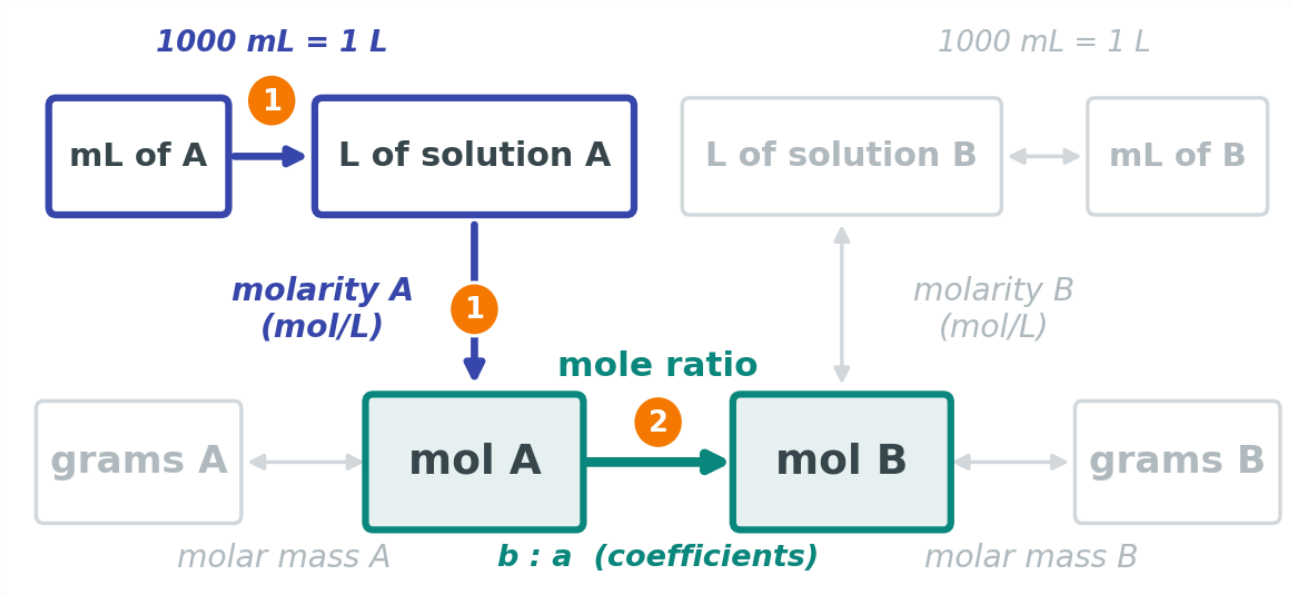
$$0.2500 \text{ L soln} \times \frac{0.400 \text{ mol HCl}}{1 \text{ L soln}} \times \frac{1 \text{ mol H}_2}{2 \text{ mol HCl}} = 0.0500 \text{ mol H}_2$$

 **Sense check:** Each H₂ consumes two HCl, so the mole count halves: 0.100 mol HCl → 0.0500 mol H₂ ✓

✓ Worked example 1: the route on the map



given: 250.0 mL of 0.400 M HCl · found: 0.0500 mol H₂



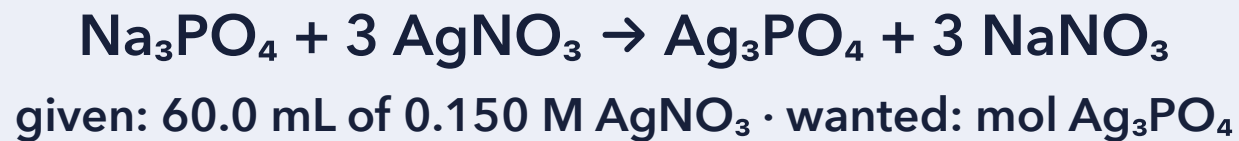
🧠 **Sense check:** A is HCl, B is H₂. Moles were wanted, so the route stops at mol B after the mole ratio. ✓

 Practice 2

Excess sodium phosphate precipitates the silver in **60.0 mL of 0.150 M AgNO₃** as yellow Ag₃PO₄. How many moles of Ag₃PO₄ form?

- A. 0.00900
- B. 0.0270
- C. 0.00300
- D. 3.00

✓ Practice 2: answer C



$$0.0600 \text{ L soln} \times \frac{0.150 \text{ mol AgNO}_3}{1 \text{ L soln}} \times \frac{1 \text{ mol Ag}_3\text{PO}_4}{3 \text{ mol AgNO}_3} = 0.00300 \text{ mol Ag}_3\text{PO}_4 \text{ (answer C)}$$

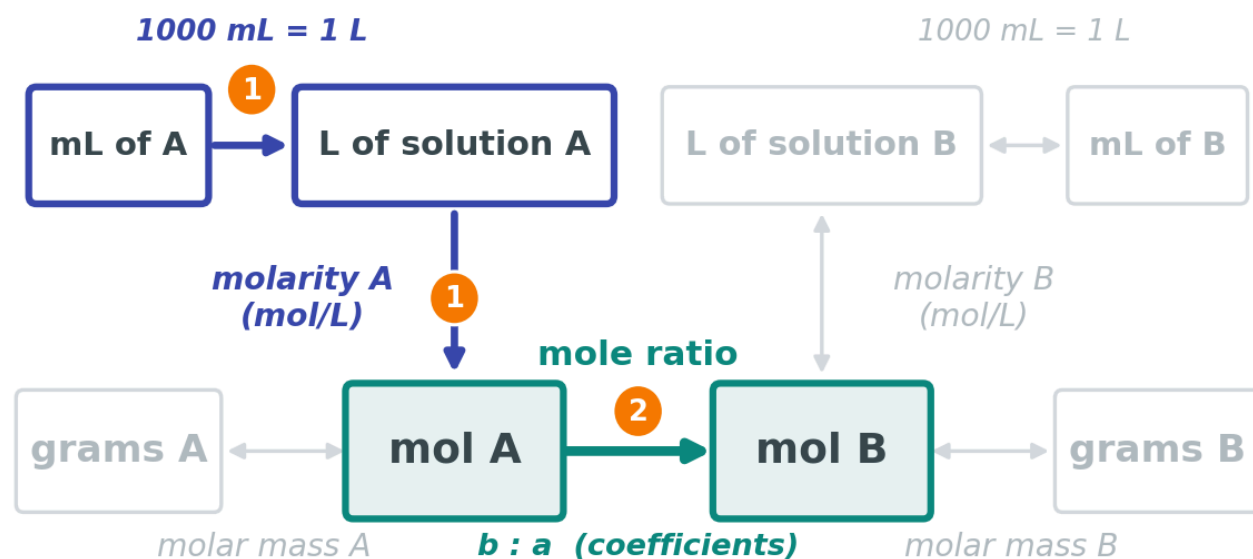
A skipped the mole ratio: $0.0600 \times 0.150 = 0.00900$ mol is the AgNO_3 . B inverted the ratio: $0.00900 \times 3 = 0.0270$. D fed milliliters to the molarity: $60.0 \times 0.150 \div 3 = 3.00$.

🧠 **Sense check:** Three AgNO_3 make one Ag_3PO_4 , so the moles drop to a third: $0.00900 \rightarrow 0.00300$ mol ✓

✓ Practice 2: the route on the map



given: 60.0 mL of 0.150 M AgNO_3 · found: 0.00300 mol Ag_3PO_4



🧠 **Sense check:** A is AgNO_3 , B is Ag_3PO_4 . A stopped at mol A and never crossed arrow 2; B crossed it upside down; D skipped the mL $\rightarrow$ L hop. ✓

Worked example 2: grams of product from a solution volume



Two clear solutions mix, and bright yellow lead(II) iodide settles out. **150.0 mL of 0.300 M KI** reacts completely with excess Pb(NO₃)₂. What mass of PbI₂ forms? (*PbI₂ 461.0 g/mol*)

A tempting shortcut: carry the moles of KI straight to grams of PbI₂. Test it against the equation.

✓ Worked example 2: solution



given: 150.0 mL of 0.300 M KI · wanted: g PbI₂ (461.0 g/mol)

Step 1 · Volume → moles

The tempting shortcut

$$0.0450 \text{ mol KI} \times \frac{461.0 \text{ g PbI}_2}{1 \text{ mol PbI}_2} = 20.7 \text{ g} \times$$

Step 2 · Moles → moles

Step 3 · Moles → the wanted unit

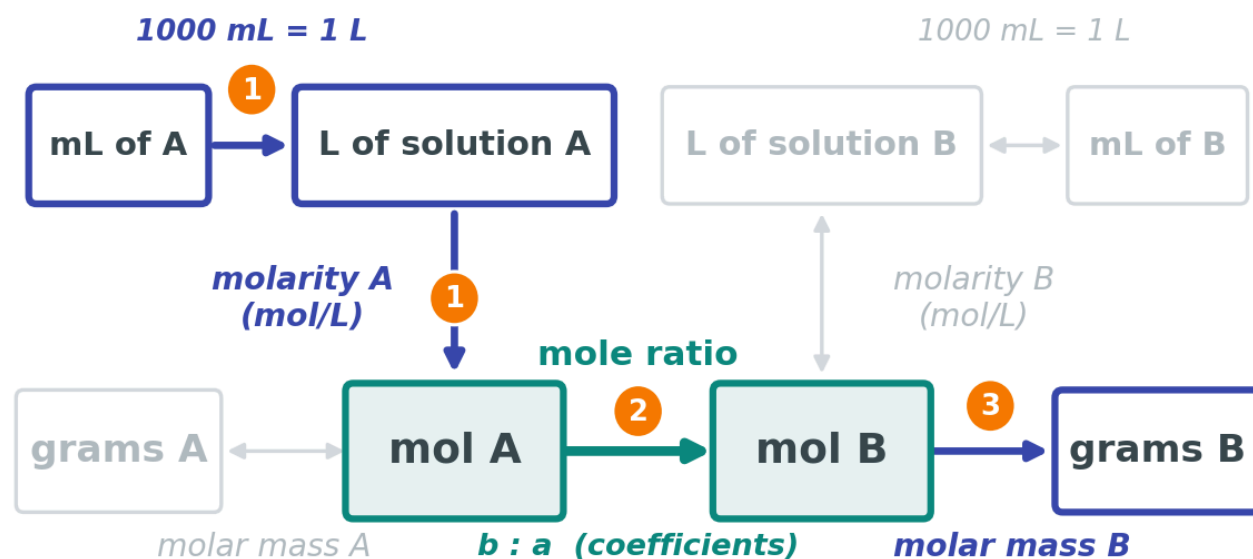
$$0.1500 \text{ L soln} \times \frac{0.300 \text{ mol KI}}{1 \text{ L soln}} \times \frac{1 \text{ mol PbI}_2}{2 \text{ mol KI}} \times \frac{461.0 \text{ g PbI}_2}{1 \text{ mol PbI}_2} = 10.4 \text{ g PbI}_2$$

 **Sense check:** Two KI deliver one PbI₂: 0.0450 mol halves to 0.0225 mol, and each mole weighs 461.0 g: about 10 g. 10.4 g ✓

✓ Worked example 2: the route on the map



given: 150.0 mL of 0.300 M KI · found: 10.4 g PbI_2



🧠 **Sense check:** A is KI, B is PbI_2 . The tempting shortcut jumped from mol A to grams B, but no arrow joins them: every route crosses the mole ratio. ✓

Your turn: copper(II) hydroxide



given: 200.0 mL (0.2000 L) of 0.250 M NaOH · wanted: g Cu(OH)₂ (97.57 g/mol)

$$0.2000 \text{ L soln} \times \frac{\text{mol NaOH}}{1 \text{ L soln}} \times \frac{\text{mol Cu(OH)}_2}{\text{mol NaOH}} \times \frac{\text{g Cu(OH)}_2}{1 \text{ mol Cu(OH)}_2}$$

$$= \text{ g Cu(OH)}_2$$

$$0.2000 \text{ L soln} \times \frac{0.250 \text{ mol NaOH}}{1 \text{ L soln}} \times \frac{1 \text{ mol Cu(OH)}_2}{2 \text{ mol NaOH}} \times \frac{97.57 \text{ g Cu(OH)}_2}{1 \text{ mol Cu(OH)}_2}$$

$$= 2.44 \text{ g Cu(OH)}_2$$

Where this goes wrong



given: 150.0 mL of 0.300 M KI · correct: 0.1500 L → 0.0450 mol KI → 0.0225 mol PbI₂ → 10.4 g

Feeding milliliters to the molarity. $150.0 \times 0.300 = 45.0$ mol KI, a thousand times too many. Molarity counts liters of solution: 150.0 mL is 0.1500 L, and $0.1500 \times 0.300 = 0.0450$ mol.

Skipping the mole ratio. $0.0450 \times 461.0 = 20.7$ g assumes one PbI₂ per KI. The equation gives 1 PbI₂ : 2 KI, and only the mole ratio switches substances.

Inverting the ratio. (2 mol KI / 1 mol PbI₂) leaves mol KI uncanceled, and $0.0450 \times 2 \times 461.0 = 41.5$ g is wrong. Write the wanted substance on top, so the given unit cancels.

 Practice 3

molar mass Fe(OH)_3 106.87 g/mol

300.0 mL of 0.200 M NaOH reacts completely with excess iron(III) chloride, and rust-brown Fe(OH)_3 precipitates. What mass of Fe(OH)_3 , in grams, forms?

- A. 2.14
- B. 6.41
- C. 19.2
- D. 0.0200

✓ Practice 3: answer A



given: 300.0 mL of 0.200 M NaOH · wanted: g Fe(OH)₃ (106.87 g/mol)

$$0.3000 \text{ L soln} \times \frac{0.200 \text{ mol NaOH}}{1 \text{ L soln}} \times \frac{1 \text{ mol Fe(OH)}_3}{3 \text{ mol NaOH}} \times \frac{106.87 \text{ g Fe(OH)}_3}{1 \text{ mol Fe(OH)}_3}$$

= 2.14 g Fe(OH)₃ (answer A)

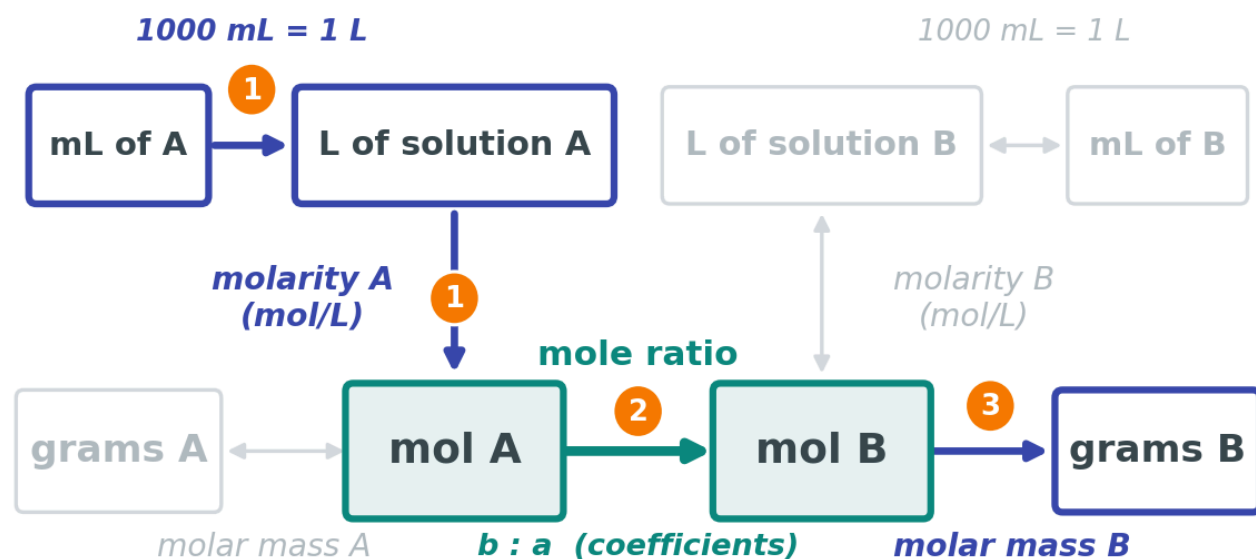
B skipped the mole ratio: $0.0600 \times 106.87 = 6.41$. C inverted the ratio: $0.0600 \times 3 \times 106.87 = 19.2$. D stopped at moles: $0.0600 \times (1/3) = 0.0200 \text{ mol Fe(OH)}_3$, one factor short of grams.

 **Sense check:** 0.0600 mol NaOH gives a third as many moles of Fe(OH)₃: 0.0200 mol, at about 107 g per mole: about 2 g. 2.14 g ✓

✓ Practice 3: the route on the map



given: 300.0 mL of 0.200 M NaOH · found: 2.14 g Fe(OH)₃



🧠 **Sense check:** A is NaOH, B is Fe(OH)₃. B skipped arrow 2, the mole ratio; C ran it upside down; D stopped at mol B, one arrow short of grams. ✓



Worked example 3: volume of a second solution



given: 175.0 mL of 0.240 M NaOH · wanted: mL of 0.150 M H₂SO₄

A base spill of **175.0 mL of 0.240 M NaOH** is neutralized with **0.150 M H₂SO₄** from the shelf. What volume of the acid solution, in milliliters, reacts completely?

Count the factors on the route: mL → L → mol NaOH → mol H₂SO₄ → L → mL.

✓ Worked example 3: solution



Step 1 • Volume → moles

$$0.1750 \text{ L NaOH soln} \times \frac{0.240 \text{ mol NaOH}}{1 \text{ L NaOH soln}} = 0.0420 \text{ mol NaOH}$$

Step 2 • Moles → moles

Step 3 • Moles → the wanted unit

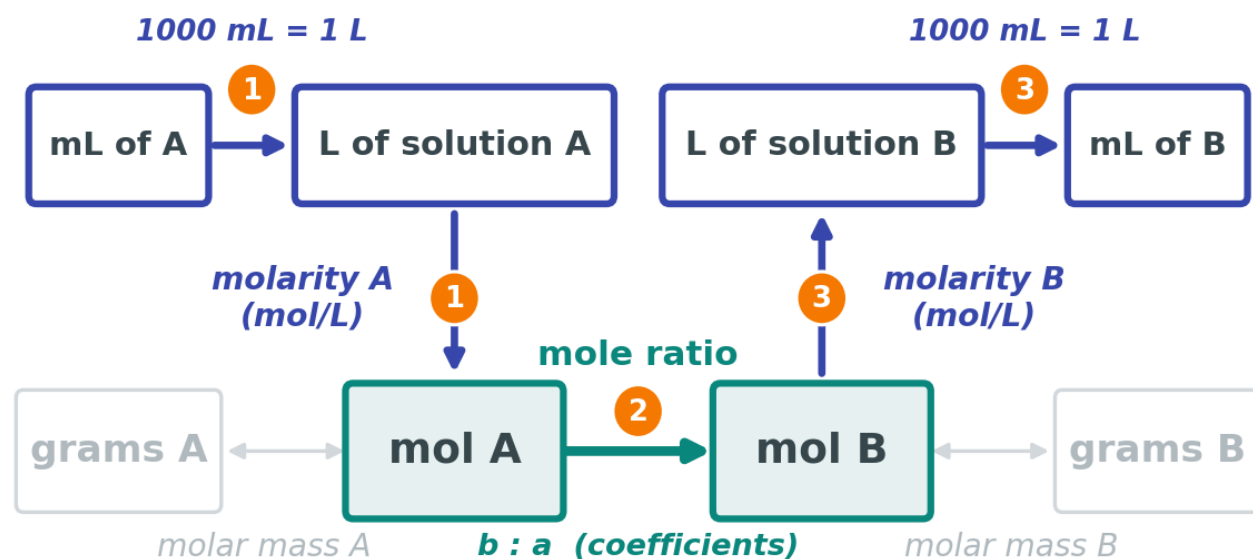
$$0.1750 \text{ L NaOH soln} \times \frac{0.240 \text{ mol NaOH}}{1 \text{ L NaOH soln}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} \times \frac{1 \text{ L acid soln}}{0.150 \text{ mol H}_2\text{SO}_4} = 0.140 \text{ L} = \mathbf{140. \text{ mL acid soln}}$$

🧠 **Sense check:** Half the moles (0.0210 vs 0.0420), but fewer per liter (0.150 vs 0.240): 140. mL is near 175.0 mL ✓

✓ Worked example 3: the route on the map



given: 175.0 mL of 0.240 M NaOH · found: 140. mL of 0.150 M H₂SO₄



🧠 **Sense check:** A is NaOH, B is H₂SO₄. Molarity works at both ends: the base's in, before the mole ratio, and the acid's out, after it. ✓



Practice 4



molar mass Mg(OH)_2 58.32 g/mol

An antacid tablet holds **0.450 g of Mg(OH)_2** . What volume of stomach acid, **0.120 M HCl**, in milliliters, does the tablet neutralize?

- A. 0.0154
- B. 64.3
- C. 32.2
- D. 0.129
- E. 129

✓ Practice 4: answer E



given: 0.450 g Mg(OH)₂ (58.32 g/mol) · wanted: mL of 0.120 M HCl

$$0.450 \text{ g Mg(OH)}_2 \times \frac{1 \text{ mol Mg(OH)}_2}{58.32 \text{ g Mg(OH)}_2} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Mg(OH)}_2} \times \frac{1 \text{ L soln}}{0.120 \text{ mol HCl}} \\ = 0.129 \text{ L} = \mathbf{129 \text{ mL (answer E)}}$$

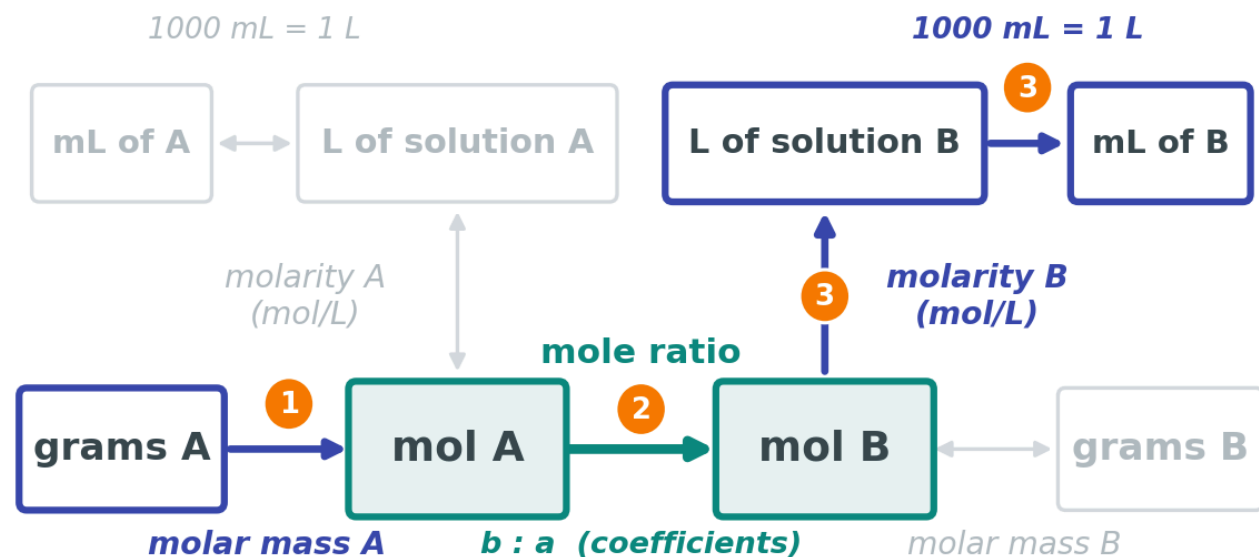
A stopped at moles: $0.450 \div 58.32 \times 2 = 0.0154 \text{ mol HCl}$. B skipped the mole ratio: $0.00772 \div 0.120 = 0.0643 \text{ L}$, or 64.3 mL. C inverted the ratio: $0.00772 \div 2 \div 0.120 = 0.0322 \text{ L}$, or 32.2 mL. D stopped at liters: 0.129 L, not yet milliliters.

 **Sense check:** Each Mg(OH)₂ takes two HCl: 0.00772 mol of base needs 0.0154 mol of acid, and at 0.120 mol per liter that is about an eighth of a liter. 129 mL ✓

✓ Practice 4: the route on the map



given: 0.450 g $\text{Mg}(\text{OH})_2$ · found: 129 mL of 0.120 M HCl



🧠 **Sense check:** A is $\text{Mg}(\text{OH})_2$, B is HCl. A mass enters through the molar mass, and the volume leaves through the acid's molarity. A stopped at mol B; D stopped at L of solution B. ✓



Practice 5



two labeled solutions: 0.250 M Ba(OH)_2 · 0.400 M HCl

150.0 mL of 0.250 M Ba(OH)_2 reacts completely with **0.400 M HCl**. What volume of the HCl solution, in milliliters, is required?

- A. 0.0750
- B. 46.9
- C. 93.8
- D. 188

✓ Practice 5: answer D



given: 150.0 mL of 0.250 M Ba(OH)₂ · wanted: mL of 0.400 M HCl

$$0.1500 \text{ L base soln} \times \frac{0.250 \text{ mol Ba(OH)}_2}{1 \text{ L base soln}} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Ba(OH)}_2} \times \frac{1 \text{ L acid soln}}{0.400 \text{ mol HCl}}$$

= 0.1875 L = **188 mL (answer D)**

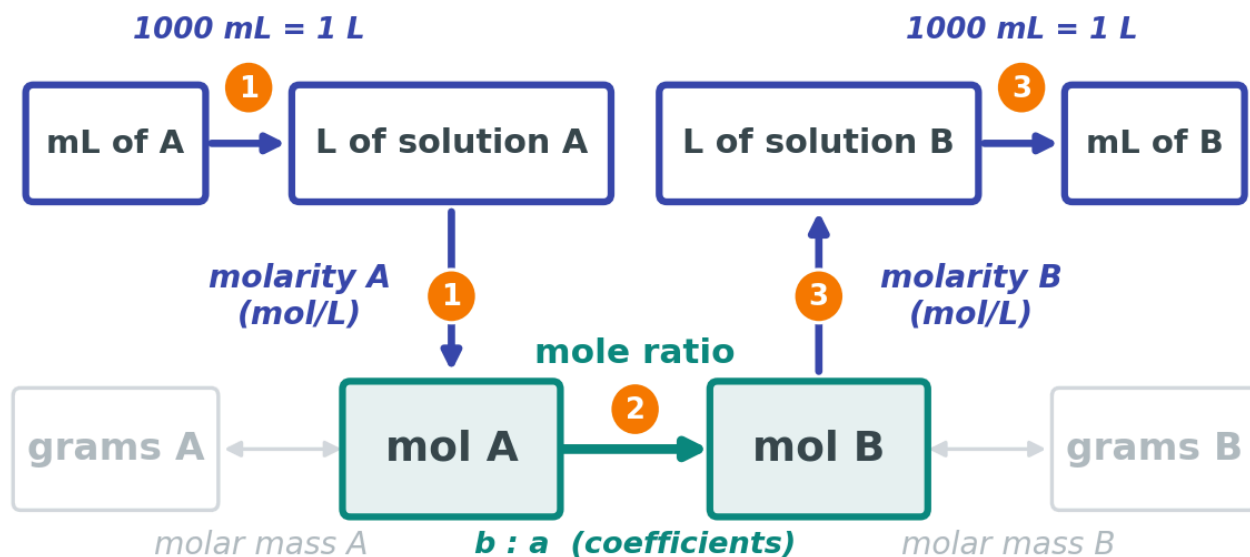
A stopped at moles: $0.0375 \times (2/1) = 0.0750 \text{ mol HCl}$, one factor short of a volume. B inverted the mole ratio: $0.0375 \times 1/2 \div 0.400 = 0.0469 \text{ L}$, or 46.9 mL. C skipped the mole ratio: $0.0375 \div 0.400 = 0.0938 \text{ L}$, or 93.8 mL.

 **Sense check:** Two HCl are needed per Ba(OH)₂, and the acid is not twice as concentrated (0.400 vs 0.250 M), so the acid volume comes out larger: 188 mL vs 150.0 mL ✓

✓ Practice 5: the route on the map



given: 150.0 mL of 0.250 M $\text{Ba}(\text{OH})_2$ · found: 188 mL of 0.400 M HCl



🧠 **Sense check:** A is $\text{Ba}(\text{OH})_2$, B is HCl. A stopped at mol B, before the acid's molarity; B ran arrow 2 upside down; C skipped it. ✓



Check yourself

1. A bottle is labeled **2.00 M NaOH**. Write the setup that converts 50.0 mL of this solution to moles of NaOH.
2. In the route **L of A → mol A → mol B → L of B**, name the conversion factor at each arrow. Which one comes from the balanced equation?

This chain also runs backward: the volume of a known-molarity solution that reacts completely with an unknown gives the unknown's moles. That laboratory measurement is a **titration**.

Can you...?

- ☐ identify solute and solvent, write dissolving equations, and predict solubility with like dissolves like?
- ☐ explain the dissolving process through the attractions broken and formed, and why dissolving can cool or warm the solution?
- ☐ classify a solution as unsaturated, saturated, or supersaturated and predict how temperature and pressure (Henry's law) change solubility?
- ☐ calculate and use molarity, percent concentration, ppm/ppb, and molality, and dilute a solution with $M_1V_1 = M_2V_2$?
- ☐ carry a solution's volume and molarity through reaction stoichiometry?
- ☐ describe the colligative properties and calculate freezing-point depression and boiling-point elevation with $\Delta T = i \cdot K \cdot m$?

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