

Solutions

CHEM 50 · Dr. Karmach

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

- Explain "like dissolves like" and predict whether a solute dissolves in a given solvent
- Describe unsaturated, saturated, and supersaturated solutions; read a solubility curve
- Use the dilution relation $M_1V_1 = M_2V_2$ to find a concentration or volume
- Calculate molality and mole fraction, and convert among molarity, molality, and mass percent
- Classify a mixture as a solution, a colloid, or a suspension from particle size, light scattering, and settling, and name a colloid's type from its dispersed phase and medium
- Apply colligative properties (freezing-point depression, boiling-point elevation, and Raoult's law) using the van't Hoff factor
- Apply Henry's law to gas solubility and titration stoichiometry to find an unknown concentration

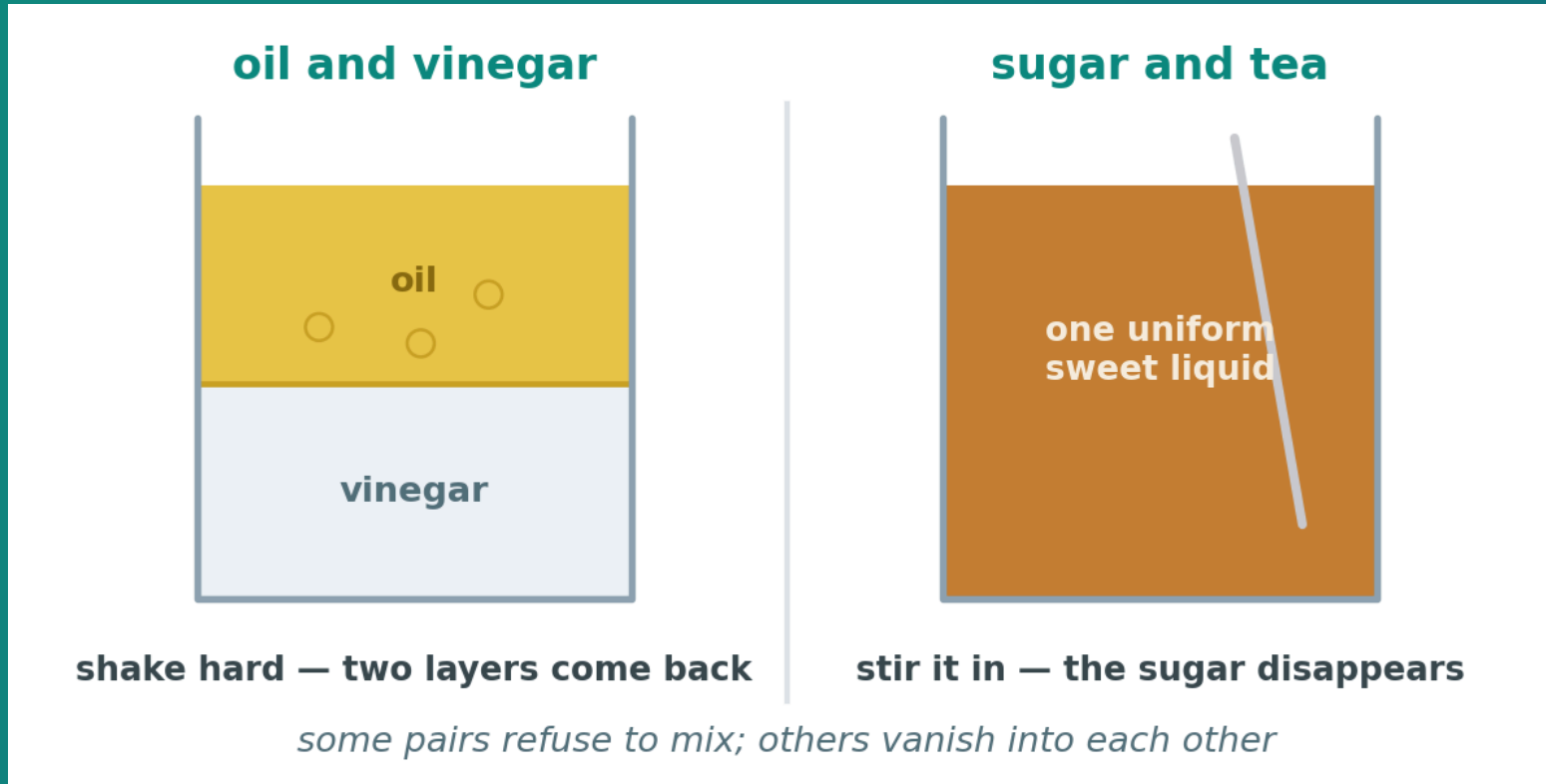
Today's route

1. Like Dissolves Like
2. Saturation & Solubility Curves
3. Dilution
4. Molality & Mole Fraction
5. Converting Concentration Units
6. Colloids
7. Freezing Point & Boiling Point
8. Raoult's Law
9. Henry's Law
10. Titration Calculations

1 • 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.

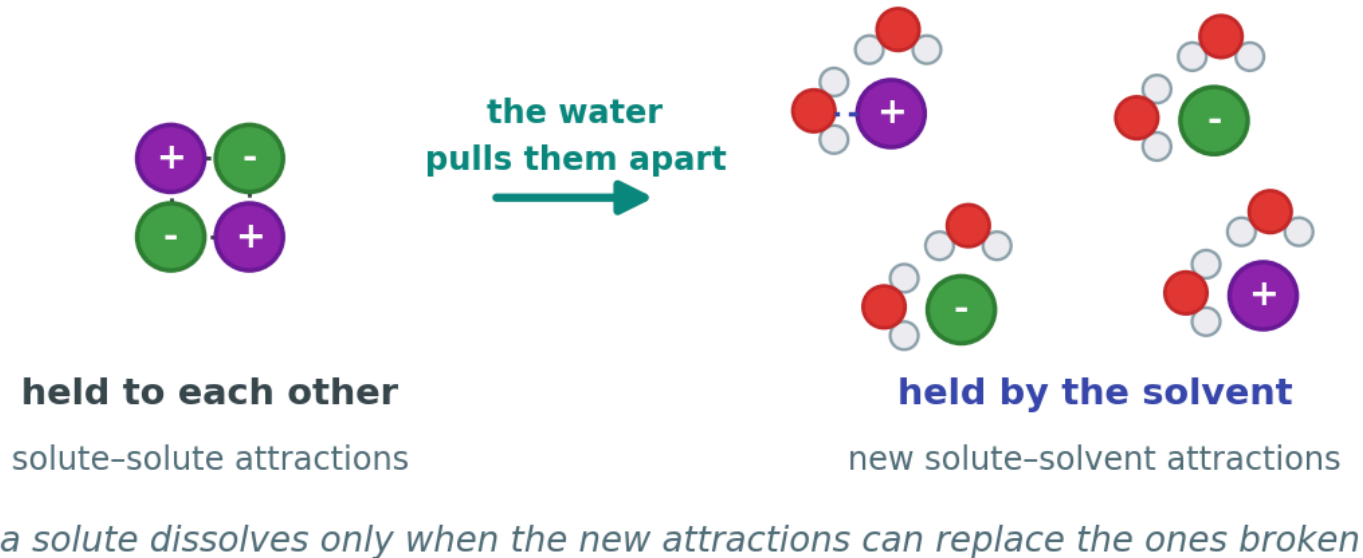
Why oil won't mix but sugar vanishes



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

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.

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

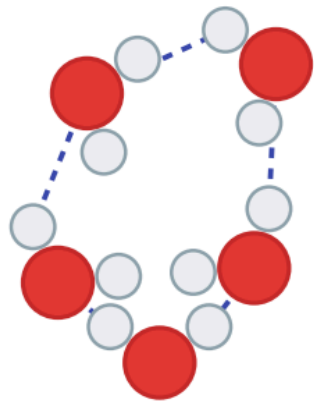
Attractions match only when polarities match. A **polar** or ionic solute dissolves in a polar solvent like water. A **nonpolar** solute dissolves in a nonpolar solvent like oil or hexane.

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

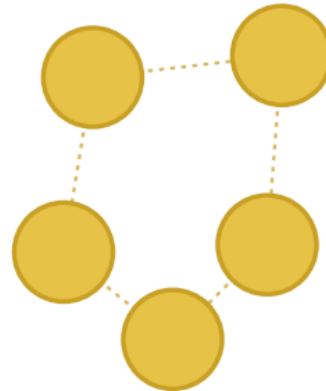
water holds itself; the oil is squeezed out



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.



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

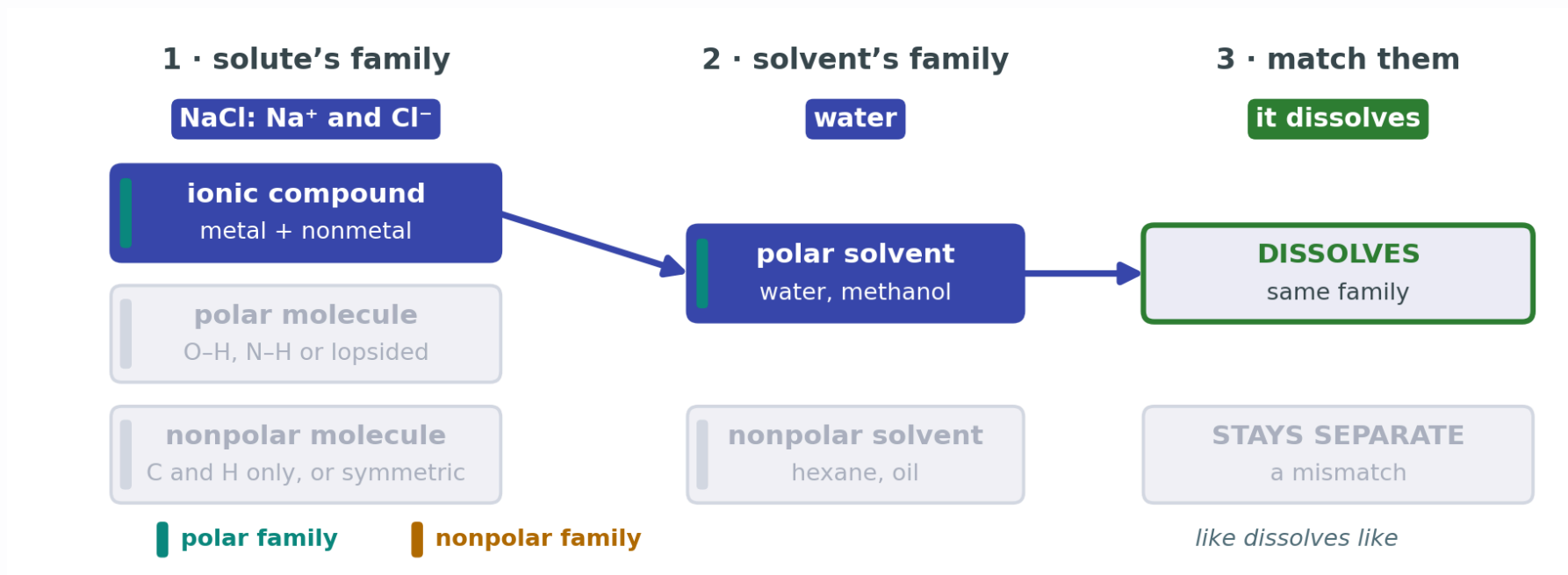
Step 3 · Match them

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

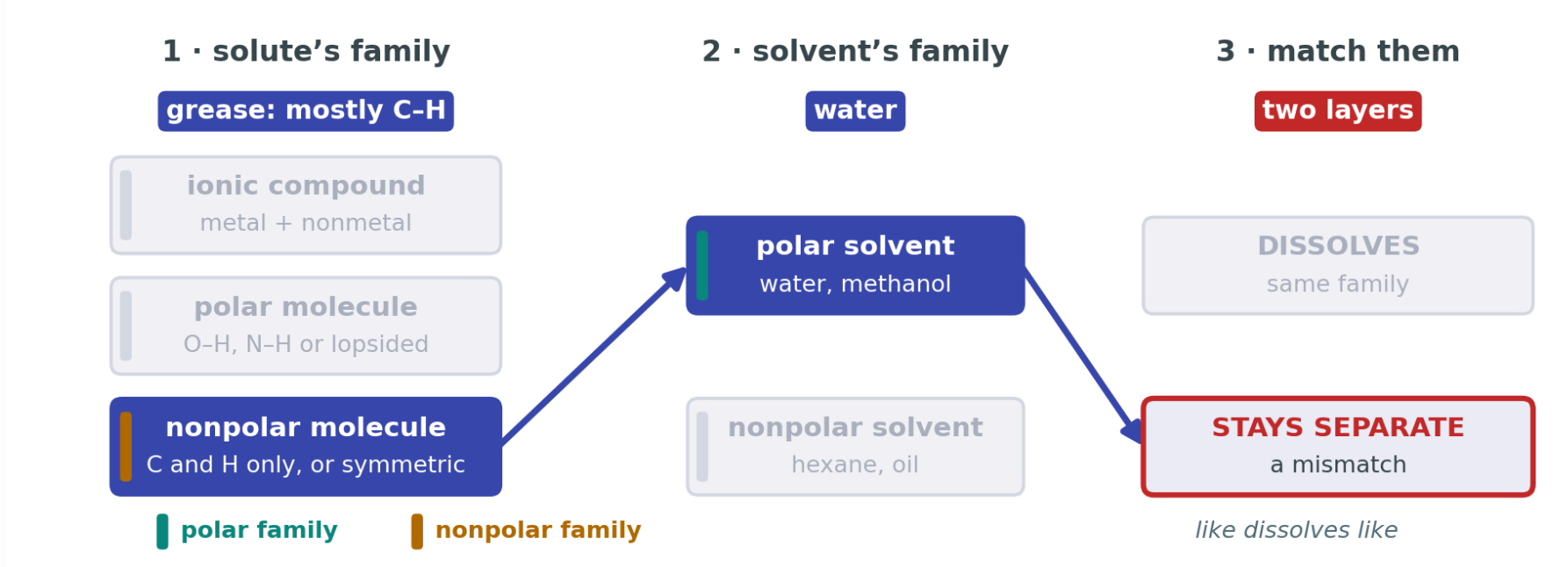
Step 3 · Match them

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



🧠 **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 1

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 1 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.



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.

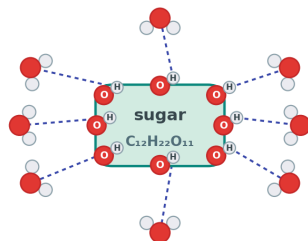
✓ Worked example 3: the polar groups

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

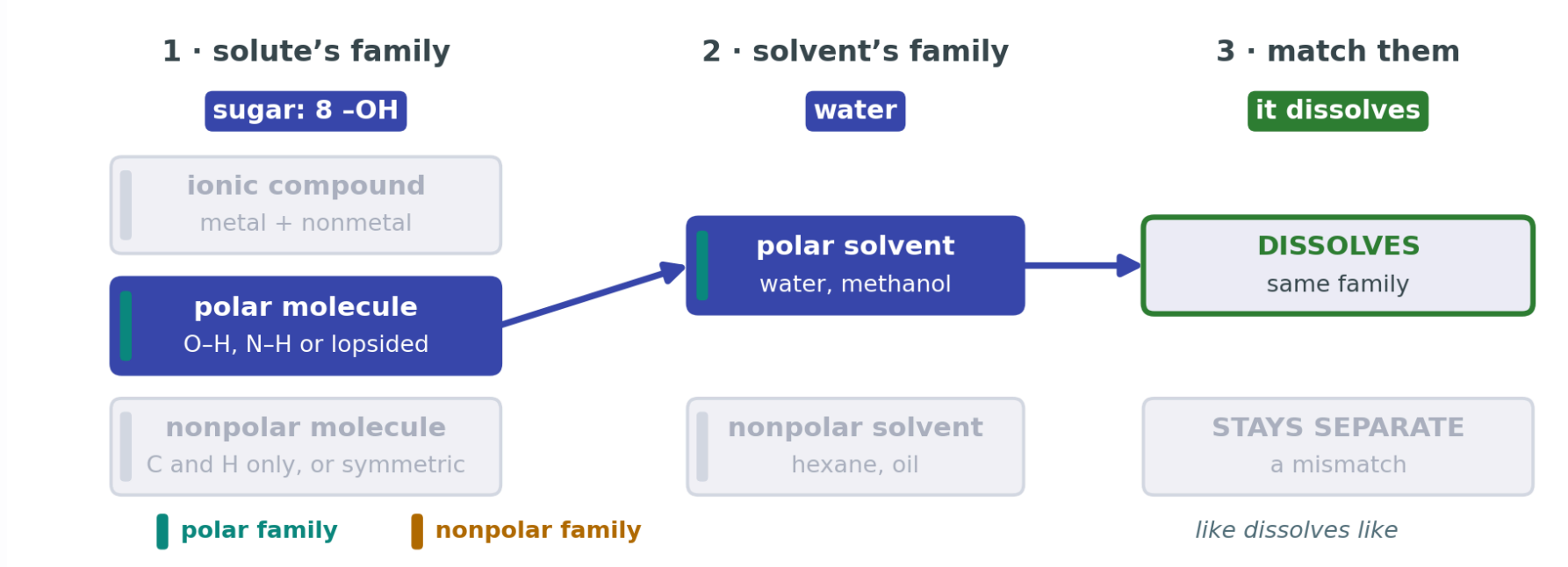
Step 3 · Match them

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



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



Practice 2

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 2 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.

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, $(\text{NH}_2)_2\text{CO}$ · anthracene, $\text{C}_{14}\text{H}_{10}$
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 anthracene
- C. None: water dissolves all three solids
- D. Anthracene only
- E. All three, once the mixture is shaken long enough

✓ Extra practice 2 answer: D

hexane: anthracene (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 ·

anthracene: only C and H $\rightarrow$ 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 treated water as a universal solvent; anthracene holds only C and H, so it has nothing polar for water to attract. E counted on shaking, which changes only how fast the solutes sort, not where they end up.

 **Sense check:** Look for ions and polar groups: KBr has ions, urea has N-H and C=O, anthracene has neither. ✓



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

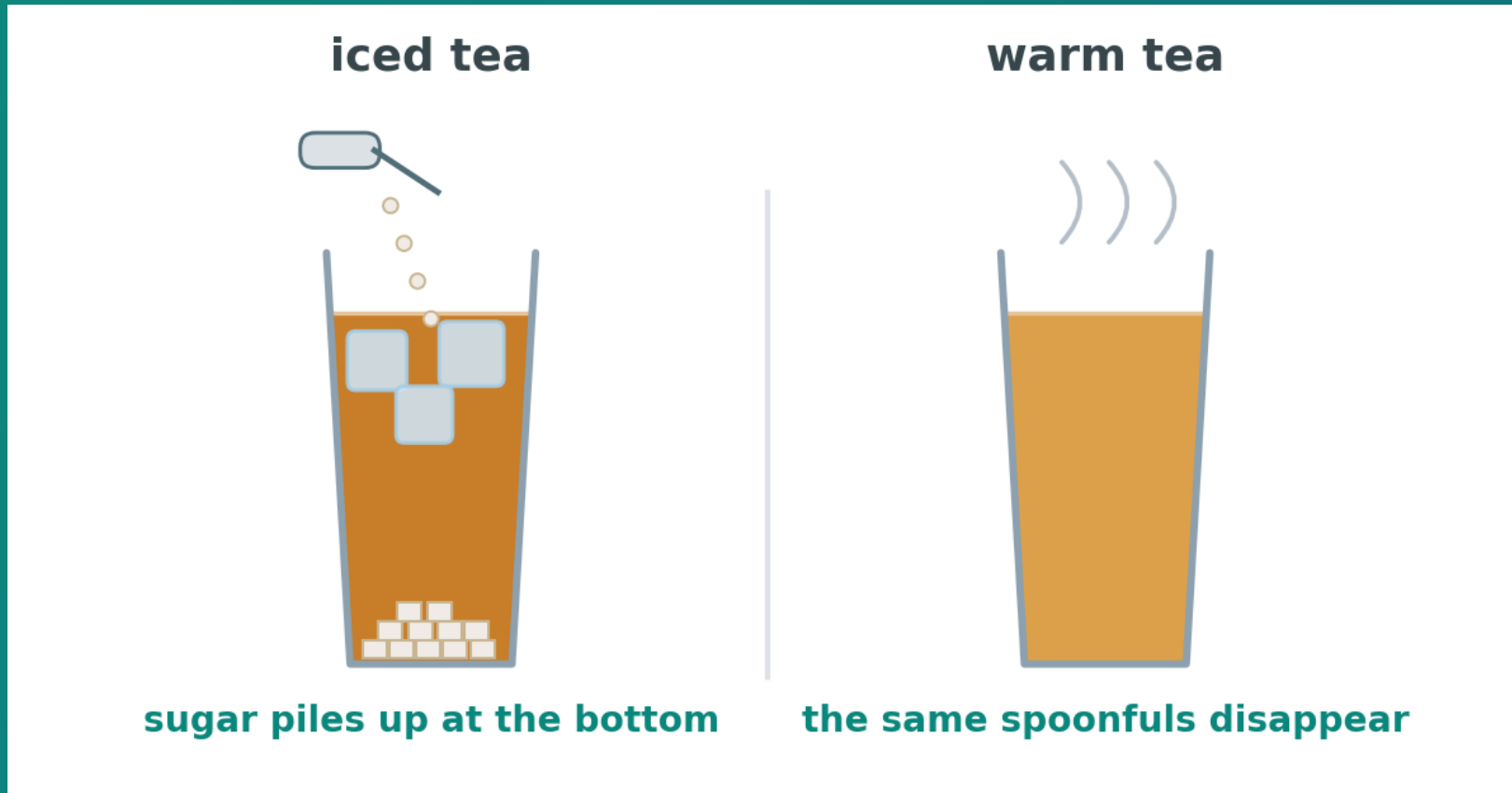
1. Iodine (I_2) is a nonpolar solid. It barely colors water but turns hexane deep purple. Which solvent dissolves it, and why?
2. Rubbing alcohol mixes with water in any amount. What does that tell you about its polarity?

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.

2 · 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

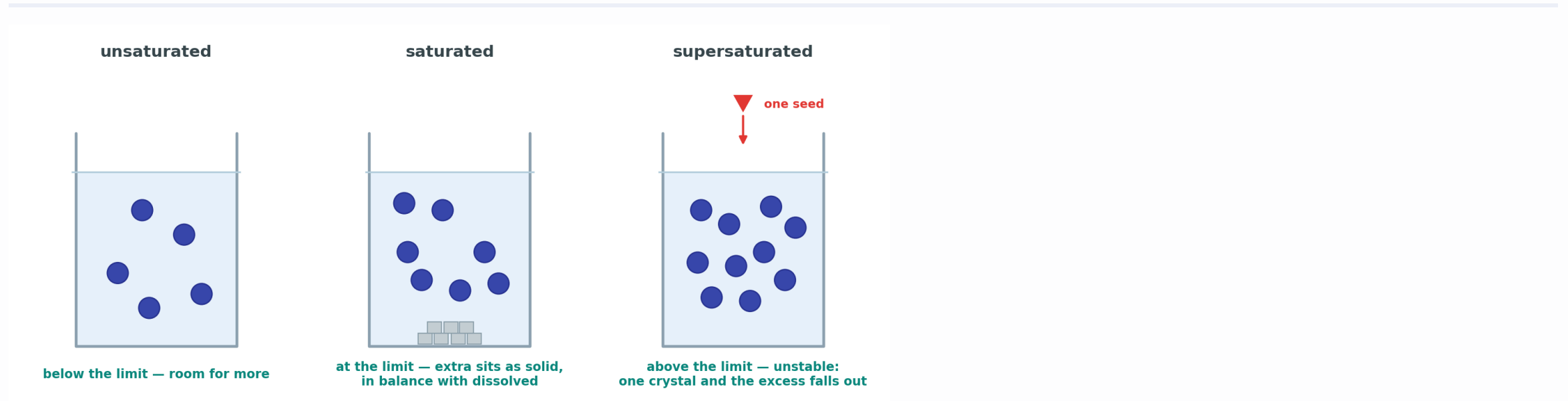
measured in grams of solute per 100 g of water, at a stated temperature

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.

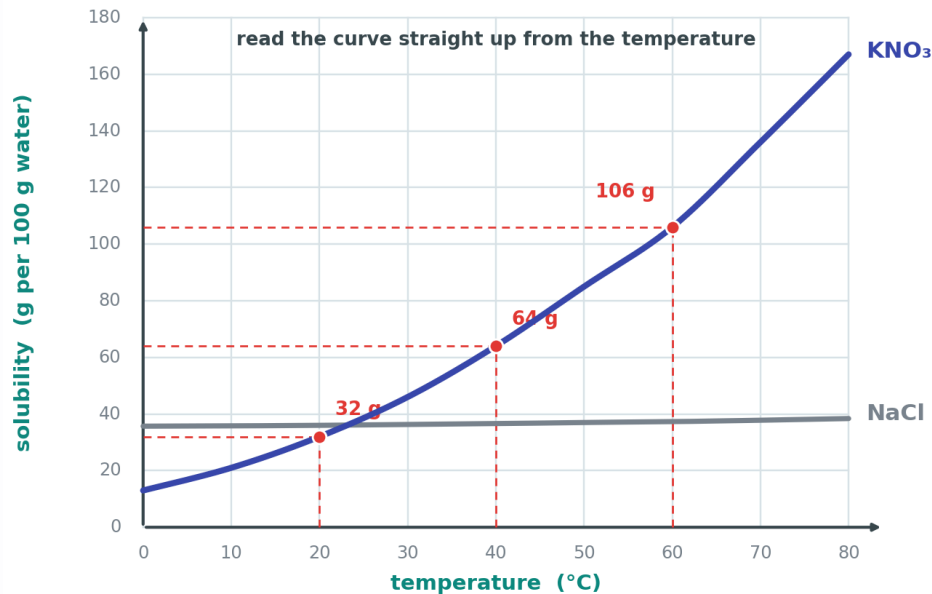
Unsaturated, saturated, supersaturated



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)

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 supersaturated.
Cooling drops the excess out.



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

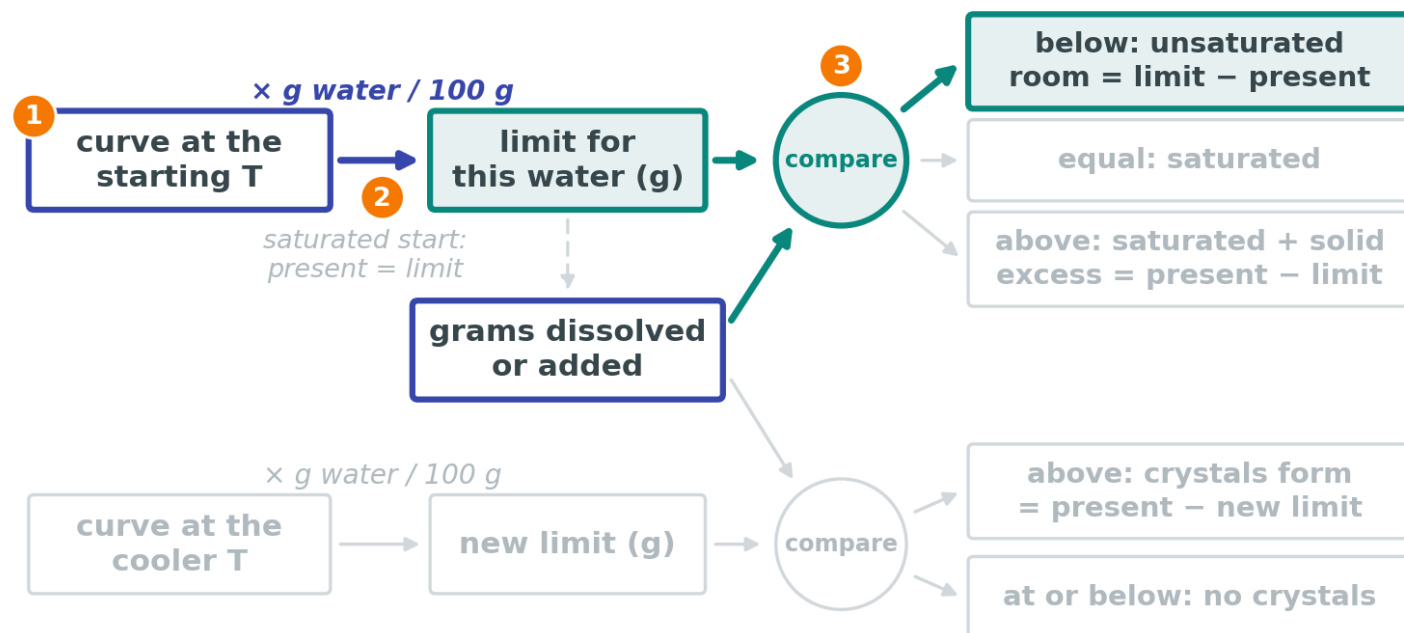
128 g limit – 100 g present = **28 g of room left** → unsaturated

 **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.



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

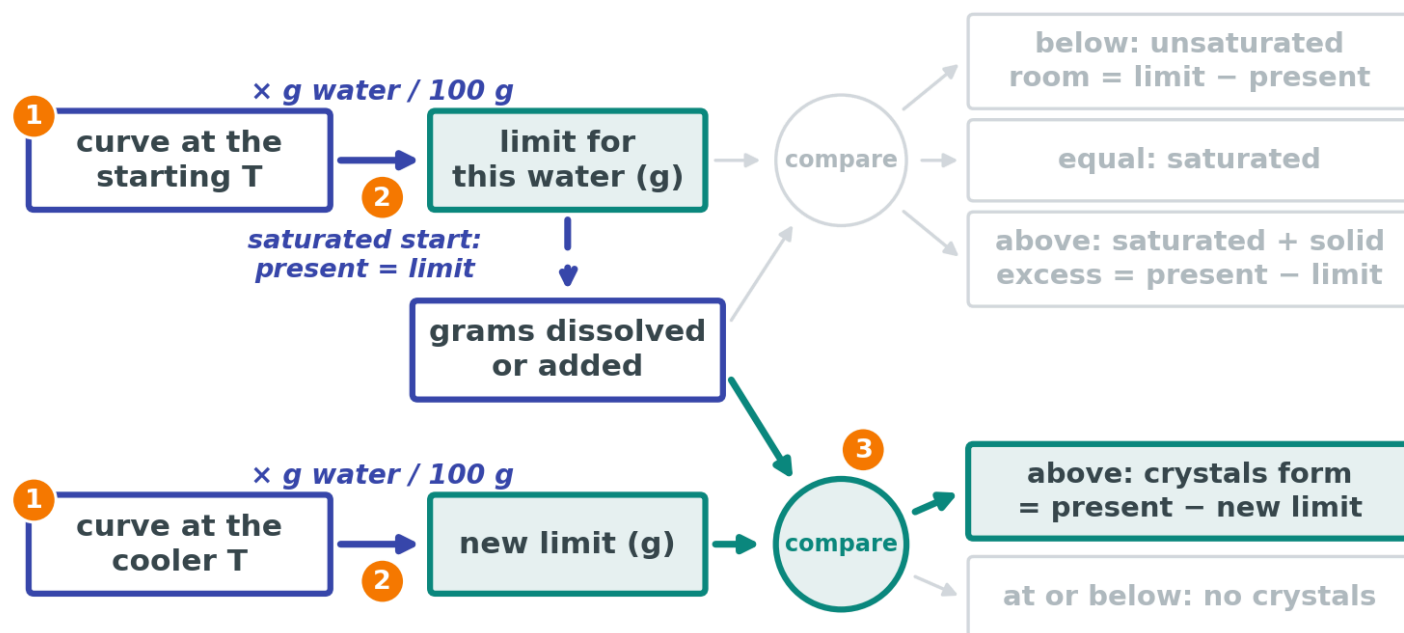
$106\text{ g dissolved} - 32\text{ g the cool water can hold} = 74\text{ g crystallize}$

 **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 · $60\text{ }^\circ\text{C} \rightarrow 20\text{ }^\circ\text{C}$

given: 100 g water · saturated at $60\text{ }^\circ\text{C}$ · cooled to $20\text{ }^\circ\text{C}$ · found: 74 g crystallize



🧠 **Sense check:** Saturated at the start, so the $60\text{ }^\circ\text{C}$ limit is the grams dissolved. The $20\text{ }^\circ\text{C}$ 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 1

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 1 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 2

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 2 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 3

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 3 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 4

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 4 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

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

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 5

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 5 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 ✓

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?

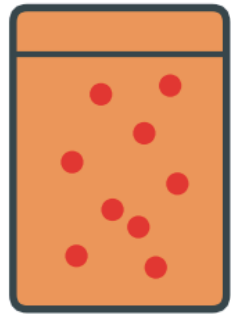
A dissolved solute raises the next question: how much of it is in the water: the **concentration**. Adding water lowers the concentration without changing the amount of solute, the relation $M_1V_1 = M_2V_2$.

3 • 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.

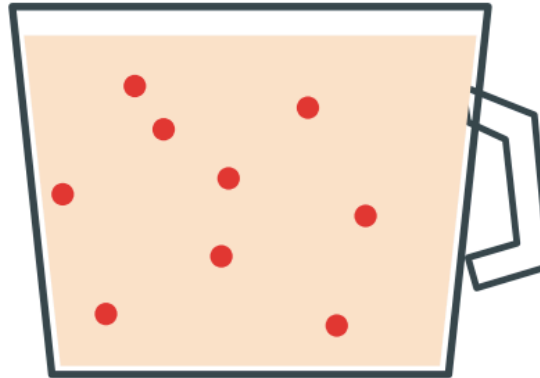
One can makes a whole pitcher

the same flavor bits — spread through much more liquid



one small can of
thick, strong concentrate

+ a pitcher of water

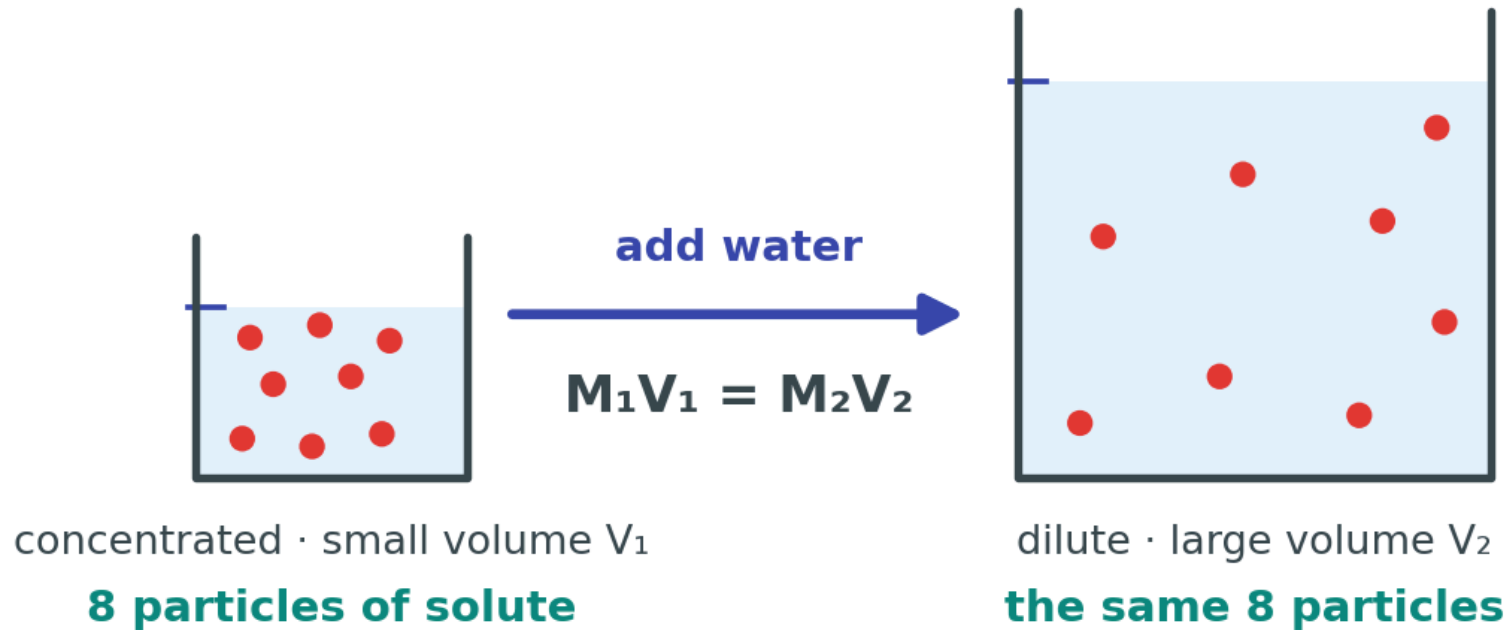


a full pitcher of
pale, mild juice

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.

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

The diagram illustrates the dilution equation $M_1V_1 = M_2V_2$. On the left side, M_1 is in a blue box with the label 'concentration before' and 'mol/L' below it. This is followed by a multiplication symbol $\times$ and V_1 in a blue box with the label 'volume before' and 'mL or L' below it. An equals sign $=$ follows. On the right side, M_2 is in a green box with the label 'concentration after' and 'mol/L' below it. This is followed by a multiplication symbol $\times$ and V_2 in a green box with the label 'volume after' and 'mL or L' below it.

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.



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 volume units:** both mL, or both L.
4. **Substitute and solve;** the diluted concentration comes out lower.



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

2 · Rearrange

3 · Match the units

4 · Substitute

if asked

M_1 V_1 M_2 V_2

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

$$M_2 = M_1 \times V_1 \div V_2$$

$$V_2 = V_1 \times M_1 \div M_2$$

$$V_1 = V_2 \times M_2 \div M_1$$

volumes

mL with mL

L → mL

solve

units cancel

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

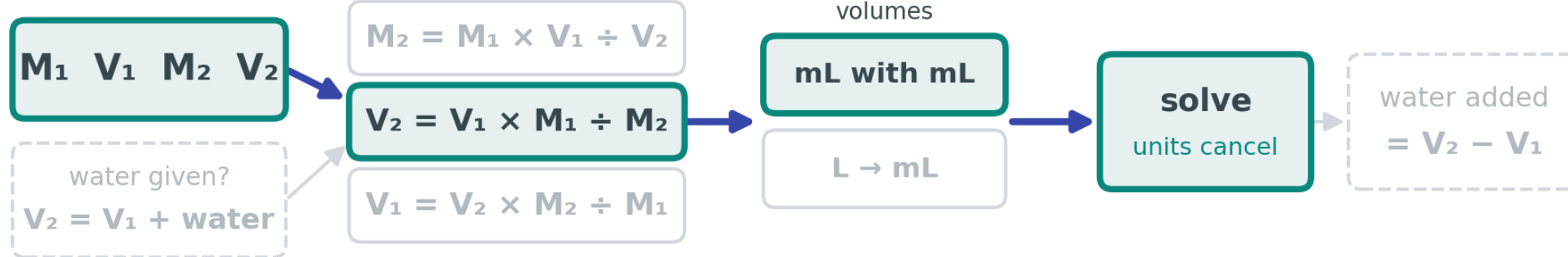
1 · List the knowns

2 · Rearrange

3 · Match the units

4 · Substitute

if asked



 **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 1

$$M_1V_1 = M_2V_2$$

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 1 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 ✓



Worked example 3: 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 3: 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 3: 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

2 · Rearrange

3 · Match the units

4 · Substitute

if asked

M_1 V_1 M_2 V_2

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

$$M_2 = M_1 \times V_1 \div V_2$$

$$V_2 = V_1 \times M_1 \div M_2$$

$$V_1 = V_2 \times M_2 \div M_1$$

volumes

mL with mL

L → mL

solve

units cancel

water added
 $= V_2 - V_1$

 **Sense check:** The unknown is the stock volume, so $V_1 = V_2 \times M_2 \div M_1$. The new move is the unit: V_2 in liters, the answer wanted in mL. ✓

 Practice 2

$$M_1V_1 = M_2V_2$$

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 2 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 ✓



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.

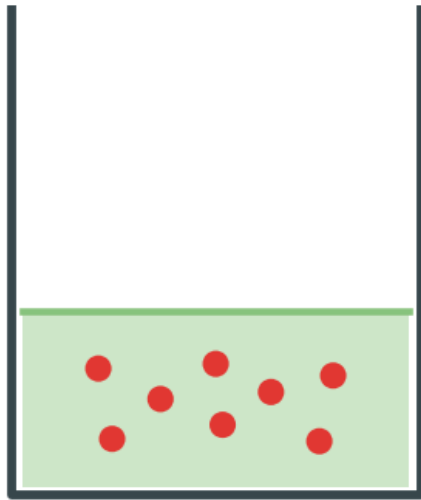
4 · Molality & Mole Fraction

Calculate the molality of a solution from grams of solute and grams of solvent, find the mole fraction of each component, and explain why these measures do not change with temperature.

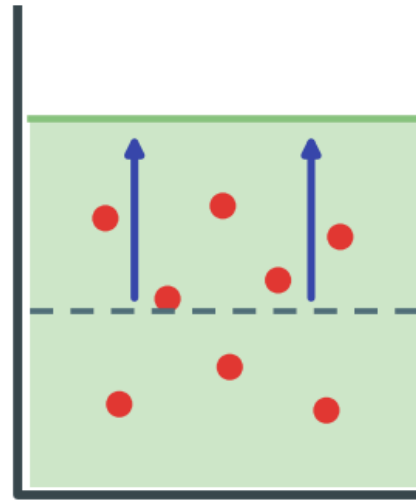
Coolant that warms and swells

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

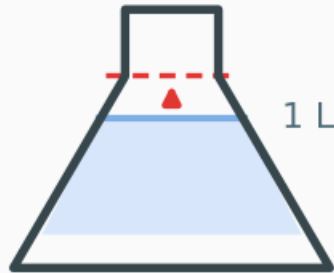
Antifreeze in a cold engine sits at a low line. Running hot, the same liquid expands and rises. Its blend never changed: only the space it fills.

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

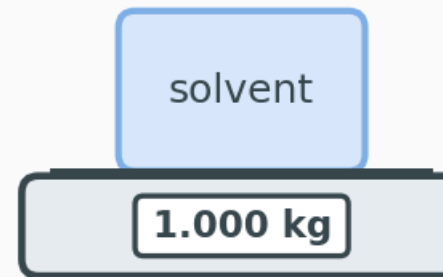


warm it — the volume rises

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 and mole fraction count moles and weigh mass, so temperature does not change them.

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

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

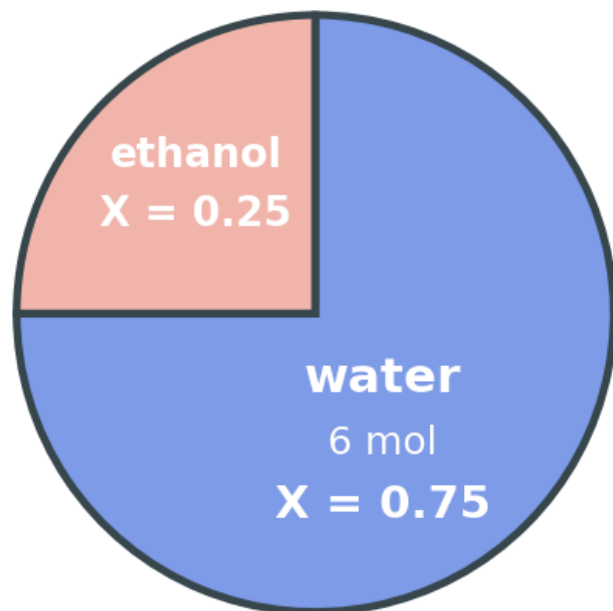
Temperature changes volume, not mass

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 and mole fraction do not.

Mole fraction: a component's share of the moles



8 mol total

a share of the total moles

$$X = \frac{\text{moles of one part}}{\text{total moles}}$$

$$\text{water: } 6 \div 8 = 0.75$$

$$\text{ethanol: } 2 \div 8 = 0.25$$

$$\mathbf{0.75 + 0.25 = 1}$$

The mole fraction X of a component is its moles divided by the total moles. It carries no units, and the fractions of all components add to 1.



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.



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

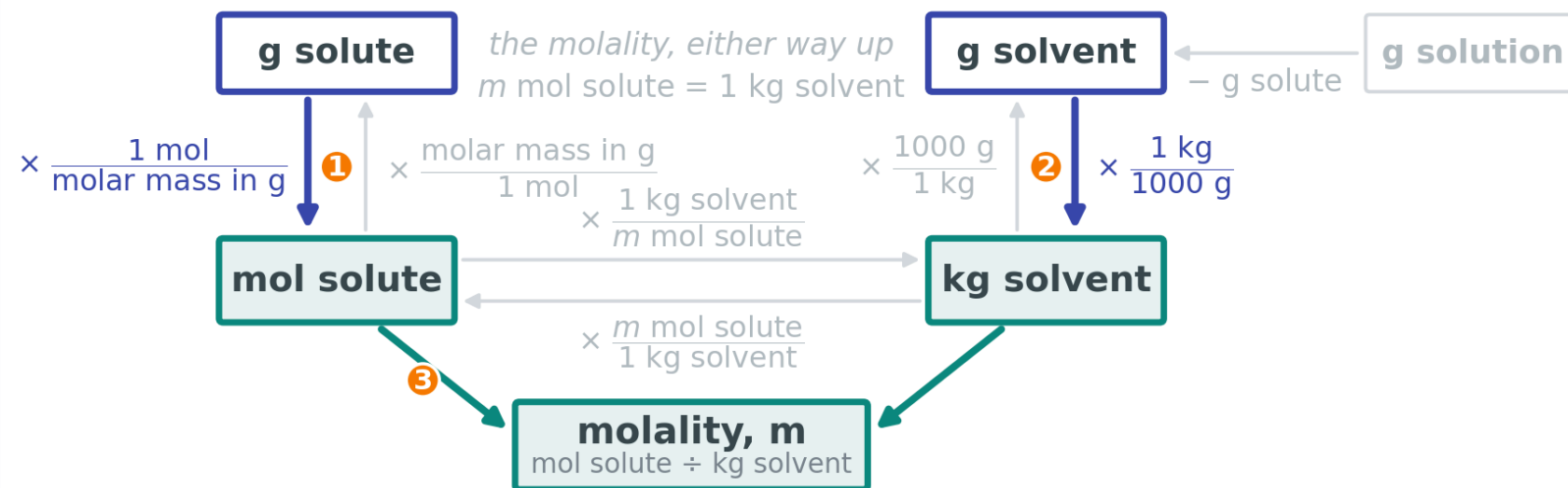
$$m = \frac{0.0999 \text{ mol glucose}}{0.250 \text{ kg water}} = 0.400 \text{ mol/kg}$$

 **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. ✓



Worked example 2: mole fraction in antifreeze

$$X = \text{mol of one component} \div \text{total mol}$$

given: 62.0 g ethylene glycol ($\text{C}_2\text{H}_6\text{O}_2$, 62.07 g/mol) · 180. g water (18.02 g/mol) · wanted: X of each

An antifreeze mix contains **62.0 g of ethylene glycol** (62.07 g/mol) and **180. g of water** (18.02 g/mol). Find the mole fraction of each component.

Set it up: both masses become moles before any fraction is taken.

✓ Worked example 2: solution

$$X = \text{mol of one component} \div \text{total mol}$$

given: 62.0 g ethylene glycol (62.07 g/mol) · 180. g water (18.02 g/mol) · wanted: X of each

Grams → moles, each component

Divide each by the total moles

$$X_{\text{glycol}} = \frac{0.999 \text{ mol}}{(0.999 + 9.99) \text{ mol}} = 0.0909$$

$$X_{\text{water}} = \frac{9.99 \text{ mol}}{10.99 \text{ mol}} = 0.909$$

 **Sense check:** The two fractions add to 1.000, and water, the greater amount, takes the larger share. $0.0909 + 0.909 = 1.000$ ✓

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**.

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

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

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

⚠️ 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: $18.0 \div 180.16 = 0.0999 \text{ mol}$.

Dividing by the solvent's moles alone. For 0.999 mol glycol and 9.99 mol water, $0.999 \div 9.99 = 0.100$. A mole fraction divides by the TOTAL: $0.999 \div 10.99 = 0.0909$.



Practice 1

$$X = \text{mol of one component} \div \text{total mol}$$

molar masses: ethanol ($\text{C}_2\text{H}_5\text{OH}$) 46.07 g/mol · water 18.02 g/mol

A mix holds **23.0 g of ethanol** and **54.0 g of water**. What is the mole fraction of ethanol?

- A. 0.143
- B. 0.167
- C. 0.299
- D. 0.499

✓ Practice 1 answer: A

$X = \text{mol of one component} \div \text{total mol}$

given: 23.0 g ethanol (46.07 g/mol) · 54.0 g water (18.02 g/mol) · wanted: X_{ethanol}

$$23.0 \text{ g} \div 46.07 \text{ g/mol} = 0.499 \text{ mol ethanol} \cdot 54.0 \text{ g} \div 18.02 \text{ g/mol} = 3.00 \text{ mol water}$$

$$X_{\text{ethanol}} = \frac{0.499 \text{ mol}}{(0.499 + 3.00) \text{ mol}} = 0.143 \cdot \text{answer A}$$

B divided by the water's moles alone: $0.499 \div 3.00 = 0.167$. C skipped the molar masses: $23.0 \div 77.0 = 0.299$ is a mass fraction. D stopped at 0.499 mol ethanol, before dividing by the total.

 **Sense check:** Water outnumbered ethanol six to one in moles, so ethanol holds about a seventh of the total: 0.143





Worked example 3: 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 3: 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

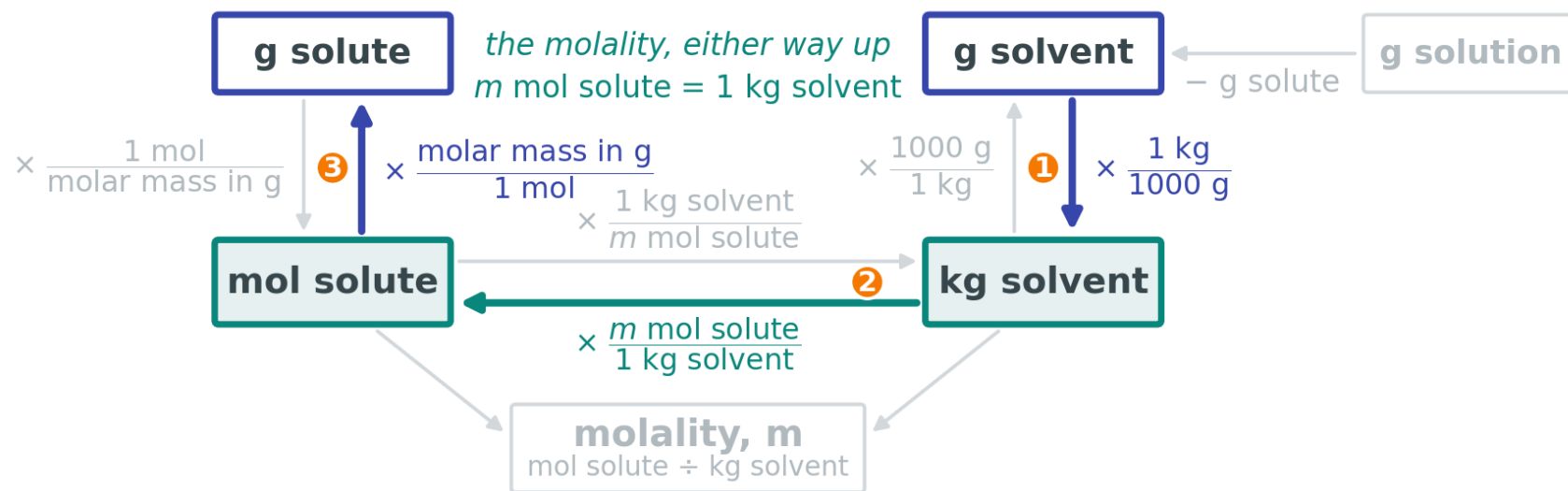
$$0.250 \text{ kg water} \times \frac{3.00 \text{ mol urea}}{1 \text{ kg water}} \times \frac{60.06 \text{ g urea}}{1 \text{ mol urea}} = 45.0 \text{ g urea}$$

 **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 3: 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:** Worked example 1's route, run 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

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

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

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 2

windshield-washer fluid: 2.00 m methanol (CH_3OH) in water
molar mass water 18.02 g/mol

A washer fluid is labeled **2.00 m methanol**. Find the methanol's mole fraction.

- A. 0.00200
- B. 0.0360
- C. 55.5
- D. 0.0348
- E. 0.965

✓ Practice 2 answer: D

2.00 m methanol: 2.00 mol methanol = 1 kg water

basis: 1 kg (1000 g) of water · 18.02 g/mol · wanted: X_{methanol}

$$1000 \text{ g water} \times \frac{1 \text{ mol water}}{18.02 \text{ g water}} = 55.5 \text{ mol water}$$

$$X_{\text{methanol}} = \frac{2.00 \text{ mol}}{(2.00 + 55.5) \text{ mol}} = 0.0348 \cdot \text{answer D}$$

A kept the water in grams: $2.00 \div 1002 = 0.00200$. B divided by the water alone: $2.00 \div 55.5 = 0.0360$. C stopped at the water's moles. E is the water's share: $55.5 \div 57.5 = 0.965$.

 **Sense check:** Methanol's molar mass never entered: the molality already counts its moles. ✓



Practice 3

pickling brine: sodium chloride (NaCl) in water

molar mass: NaCl 58.44 g/mol

How many grams of NaCl dissolve in **350. g of water** to make a **1.40 m** brine?

- A. 0.490
- B. 28.6
- C. 14.6
- D. 2.86×10^4

✓ Practice 3 answer: B

1.40 m NaCl: 1.40 mol NaCl = 1 kg water
 given: 1.40 m NaCl (58.44 g/mol) · 350. g water · wanted: g NaCl

$$0.350 \text{ kg water} \times \frac{1.40 \text{ mol NaCl}}{1 \text{ kg water}} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = \mathbf{28.6 \text{ g NaCl} \cdot \text{answer B}}$$

A stopped at moles: $0.350 \times 1.40 = 0.490 \text{ mol NaCl}$. C divided the kilograms by the molality: $0.350 \div 1.40 = 0.250$, then $\times 58.44 = 14.6$. D kept the water in grams: $350 \times 1.40 \times 58.44 = 2.86 \times 10^4$, a thousand times too much.

 **Sense check:** Less than half a kilogram of water takes less than half of 1.40 mol: 0.490 mol, and $0.490 \text{ mol} \times 58.44 \text{ g/mol} = 28.6 \text{ g}$ ✓



Practice 4

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 4 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 5

$$X = \text{mol of one component} \div \text{total mol} \cdot m = \text{mol solute} \div \text{kg solvent}$$

molar mass: water 18.02 g/mol

In an aqueous glycerol solution, the mole fraction of glycerol is **0.0600**. What is the molality of the solution?

- A. 3.33
- B. 3.54
- C. 0.0638
- D. 16.9
- E. 3.54×10^{-3}

✓ Practice 5 answer: B

$X_{\text{glycerol}} = 0.0600$: 1 mol of solution holds 0.0600 mol glycerol + 0.9400 mol water
basis: 1 mol of solution · water 18.02 g/mol · wanted: m

$$0.9400 \text{ mol water} \times \frac{18.02 \text{ g water}}{1 \text{ mol water}} = 16.94 \text{ g water} = 0.01694 \text{ kg water}$$

$$m = \frac{0.0600 \text{ mol glycerol}}{0.01694 \text{ kg water}} = 3.54 \text{ mol/kg} \cdot \text{answer B}$$

A counted the whole 1 mol as water: $0.0600 \div 0.01802 \text{ kg} = 3.33$. C divided by the water's moles: $0.0600 \div 0.9400 = 0.0638$, a mole ratio. D stopped at 16.9 g of water. E kept the water in grams: $0.0600 \div 16.94 = 3.54 \times 10^{-3}$.

 **Sense check:** 16.94 g of water carries 0.0600 mol, so a full kilogram carries 59.0 times that: 3.54 mol ✓

Check yourself

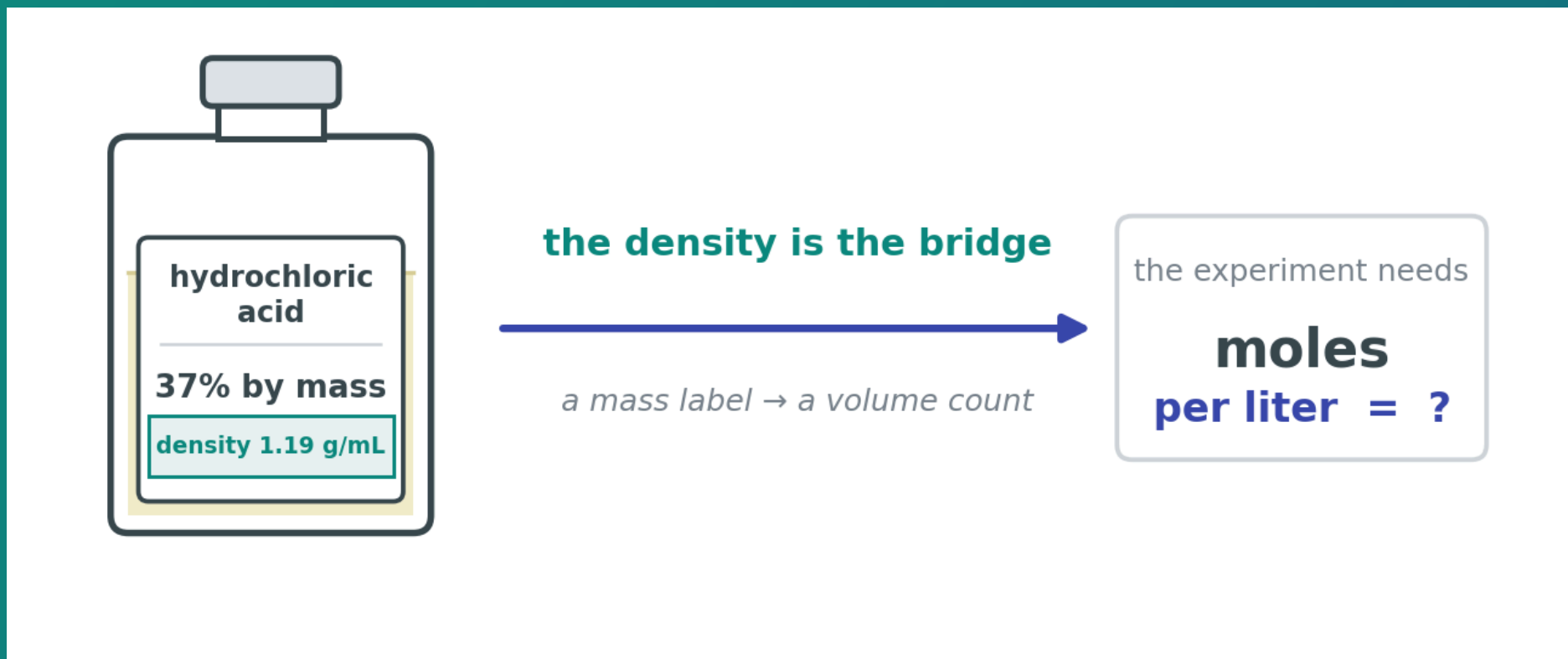
1. A solution is prepared from grams of solute and grams of water. Which mass belongs in the denominator of molality, the solvent's or the whole solution's, and in what unit?
2. A two-component mixture has $X = 0.30$ for one component. What is the mole fraction of the other?

Molarity counts moles per liter of solution; molality counts moles per kilogram of solvent. Given the solution's density, either concentration can be converted into the other.

5 · Converting Concentration Units

Convert a solution's concentration among mass percent, molarity, and molality by taking a convenient basis and bridging with density and molar mass.

The label reads 37%. The lab wants moles.



A jug of hydrochloric acid is labeled 37% by mass. The experiment needs it in moles per liter. The density on the label connects the two.

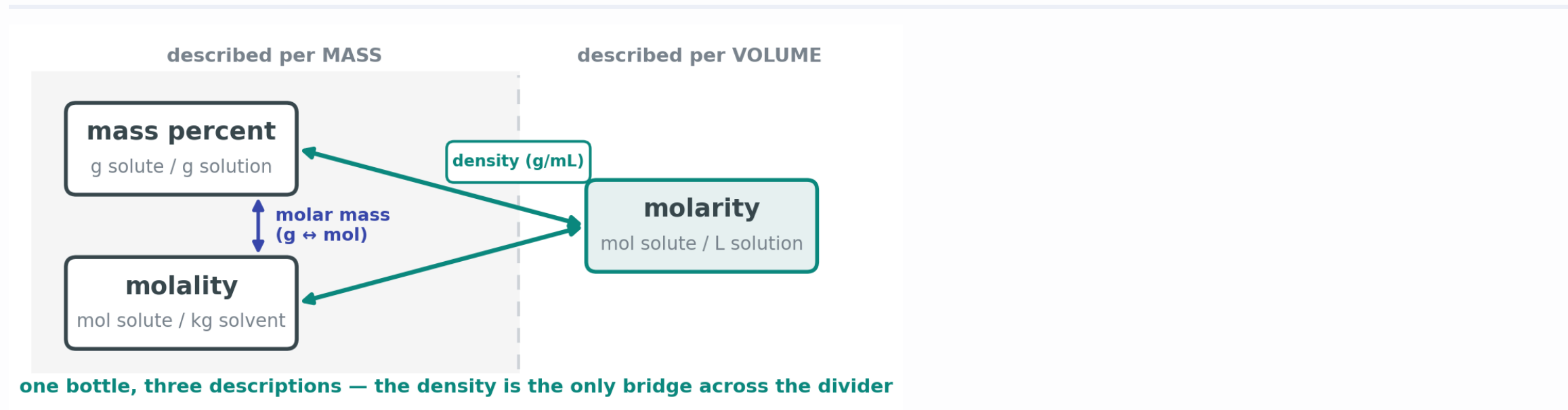
Concentration labels are conversion factors

A concentration is a ratio: solute per amount of solution or solvent. Any ratio works as a conversion factor. Density, mass percent, and molar mass are all familiar. Chained, their units cancel one after another.

$$\cancel{\text{mL soln}} \times \frac{\cancel{\text{g soln}}}{\cancel{\text{mL soln}}} \times \frac{\cancel{\text{g solute}}}{\cancel{\text{g soln}}} \times \frac{\text{mol solute}}{\cancel{\text{g solute}}} = \text{mol solute}$$

density (g/mL) · mass percent (g per 100 g) · molar mass (g/mol)
each is a ratio of two amounts · each cancels one unit and brings in the next

One solution, three descriptions



The same bottle can be reported three ways: mass percent, molality, or molarity. Each counts the solute against a different amount. Density and molar mass carry one description into another.

memory hook: M per L solution · m per kg solvent · % per 100 g solution
three denominators, one solute · solvent = solution – solute

Density and molar mass convert the units

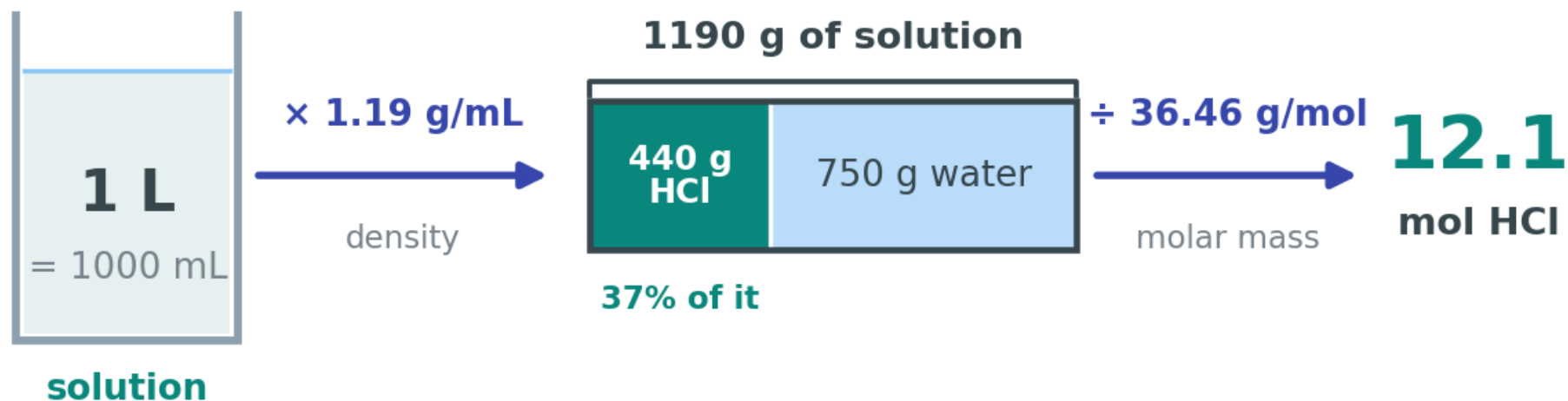
$$1 \text{ mL of solution} = (\text{density}) \text{ grams} \cdot 1 \text{ mol of solute} = (\text{molar mass}) \text{ grams}$$

density crosses volume $\leftrightarrow$ mass $\cdot$ molar mass crosses mass $\leftrightarrow$ moles

Mass percent and molality describe the solute per unit mass; molarity describes it per unit volume. Density carries a mass to a volume; molar mass carries a mass to a mole count.

Pick a convenient basis

take 1 L of 37% (m/m) HCl — density 1.19 g/mL, molar mass 36.46 g/mol

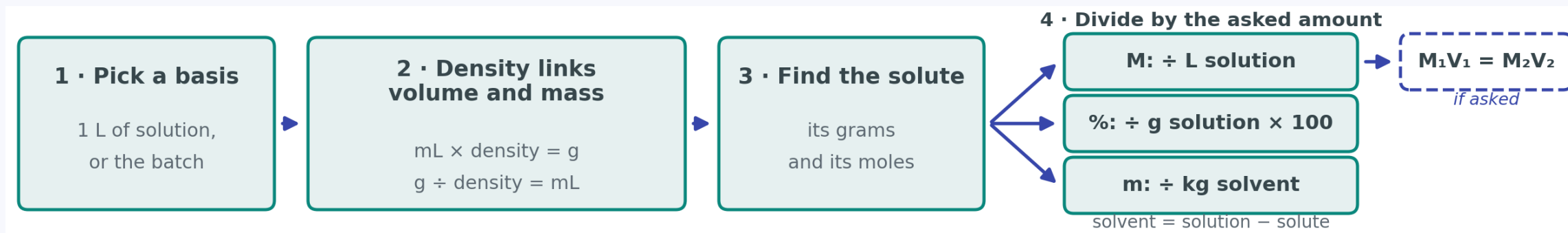


12.1 mol in the 1 L we took = 12.1 M

Concentration is a rate: the same for a drop or a drum. Take whatever amount makes the numbers easy: 1 L for molarity, 100 g for mass percent or molality. The ratio is the same.

The method

1. **Pick a basis:** 1 L, or the batch.
2. **Density links volume and mass:** $\text{mL} \times \text{density} = \text{g}$; $\text{g} \div \text{density} = \text{mL}$.
3. **Find the solute:** grams, moles.
4. **Divide by the asked amount:** L solution, g solution, kg solvent.





Worked example 1: mass percent to molarity

37.0% (m/m) HCl · density 1.19 g/mL · molar mass 36.46 g/mol

given: the label · wanted: molarity · basis: take 1 L of solution

Concentrated **hydrochloric acid** is **37.0% by mass**, with a density of **1.19 g/mL**. What is its molarity?

Take 1 L, then convert with the density and the molar mass.

✓ Worked example 1: solution

37.0% (m/m) HCl · density 1.19 g/mL · molar mass 36.46 g/mol

Step 1 · Pick a basis

Step 2 · Density links volume and mass

Step 3 · Find the solute

$$1000 \text{ mL soln} \times \frac{1.19 \text{ g soln}}{1 \text{ mL soln}} \times \frac{37.0 \text{ g HCl}}{100 \text{ g soln}} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} = 12.1 \text{ mol HCl}$$

Step 4 · Divide by the asked amount

$$M = \frac{12.1 \text{ mol HCl}}{1 \text{ L solution}} = 12.1 \text{ M}$$

🧠 **Sense check:** 12.1 mol fill the 1 L chosen: 12.1 M. The liter weighs 1190 g, over a third of it HCl. ✓

1 · Basis: 1 L soln



2 · Density



3 · Solute



4 · M: ÷ L soln



Worked example 2: concentrated sulfuric acid

98.0% (m/m) H_2SO_4 · density 1.84 g/mL · molar mass 98.08 g/mol

given: the label · wanted: molarity · basis: take 1 L of solution

Concentrated **sulfuric acid** is **98.0% by mass**, with a density of **1.84 g/mL**. What is its molarity?

A common first attempt reads 1 L as 1000 g of solution. Test that against the density.

✓ Worked example 2: solution

98.0% (m/m) H_2SO_4 · density 1.84 g/mL · molar mass 98.08 g/mol

Step 1 · Pick a basis

A common first attempt

$$1000 \text{ g soln} \times \frac{98.0 \text{ g H}_2\text{SO}_4}{100 \text{ g soln}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{98.08 \text{ g H}_2\text{SO}_4} = 9.99 \text{ M} \times$$

Step 2 · Density links volume and mass

Step 3 · Find the solute

Step 4 · Divide by the asked amount

$$1000 \text{ mL soln} \times \frac{1.84 \text{ g soln}}{1 \text{ mL soln}} \times \frac{98.0 \text{ g H}_2\text{SO}_4}{100 \text{ g soln}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{98.08 \text{ g H}_2\text{SO}_4} = 18.4 \text{ mol H}_2\text{SO}_4$$

🧠 **Sense check:** 18.4 mol in 1 L: 18.4 M. Skipping the density gave 9.99 M; the liter weighs 1840 g, not 1000. ✓

1 · Basis: 1 L soln



2 · Density



3 · Solute



4 · M: ÷ L soln

Take-home: weigh the liter with the density

$$1000 \text{ g soln} \times \frac{98.0 \text{ g H}_2\text{SO}_4}{100 \text{ g soln}} \times \frac{1 \text{ mol}}{98.08 \text{ g H}_2\text{SO}_4} = 9.99 \text{ M} \times 1 \text{ L read as 1000 g}$$

$$1000 \text{ mL soln} \times \frac{1.84 \text{ g soln}}{1 \text{ mL soln}} \times \frac{98.0 \text{ g H}_2\text{SO}_4}{100 \text{ g soln}} \times \frac{1 \text{ mol}}{98.08 \text{ g H}_2\text{SO}_4} = \mathbf{18.4 \text{ M} \checkmark}$$

The density sets the mass of a liter of solution. Skip it and every molarity comes out too low. Weigh the liter first, then take the solute's share.

Your turn: nitric acid

30.0% (m/m) HNO_3 · density 1.18 g/mL · molar mass 63.01 g/mol

basis: take 1 L of solution · wanted: molarity

A cleaning-grade nitric acid is **30.0% by mass**, density **1.18 g/mL**.

$$1000 \text{ mL soln} \times \frac{\text{g soln}}{1 \text{ mL soln}} \times \frac{30.0 \text{ g HNO}_3}{100 \text{ g soln}} \times \frac{1 \text{ mol HNO}_3}{\text{g HNO}_3} = \text{M}$$

Fill the density, the molar mass, and the molarity, then compute.

$$1000 \text{ mL soln} \times \frac{1.18 \text{ g soln}}{1 \text{ mL soln}} \times \frac{30.0 \text{ g HNO}_3}{100 \text{ g soln}} \times \frac{1 \text{ mol HNO}_3}{63.01 \text{ g HNO}_3} = 5.62 \text{ M}$$

1 · Basis: 1 L soln

2 · Density

3 · Solute

4 · M: ÷ L soln

⚠ Where this goes wrong

37.0% (m/m) HCl · density 1.19 g/mL · molar mass 36.46 g/mol

correct: 1 L → 1190 g → 440.3 g HCl → 12.1 mol → 12.1 M

Skipping the density. Reading 1 L as 1000 g: $0.370 \times 1000 \div 36.46 = 10.1$ M. But a liter of this acid weighs $1000 \times 1.19 = 1190$ g, not 1000. With the density: 12.1 M.

Dividing by the density. $0.370 \times 1000 \div 1.19 \div 36.46 = 8.53$ M. Mass is volume $\times$ density, so the milliliters multiply by 1.19. Multiplying gives 1190 g and 12.1 M.

Dropping the 1000 mL in a liter. $0.370 \times 1.19 \div 36.46 = 0.0121$: a thousand times too small. A liter is 1000 mL, so 1 L weighs 1190 g and the molarity is 12.1 M.

 Practice 1

15.0% (m/m) NaOH · density 1.16 g/mL · molar mass 40.00 g/mol

basis: take 1 L of solution · wanted: molarity

A stock **15.0% (m/m) sodium hydroxide** solution has a density of **1.16 g/mL**. What is its molarity?

- A. 4.35
- B. 3.23
- C. 3.75
- D. 0.00435

✓ Practice 1 answer: A

15.0% (m/m) NaOH · density 1.16 g/mL · molar mass 40.00 g/mol

basis: take 1 L → 1160 g solution · wanted: molarity

$$1000 \text{ mL soln} \times \frac{1.16 \text{ g soln}}{1 \text{ mL soln}} \times \frac{15.0 \text{ g NaOH}}{100 \text{ g soln}} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} = 4.35 \text{ M} \cdot \text{answer A}$$

C skipped the density: $0.150 \times 1000 \div 40.00 = 3.75$, reading 1 L as 1000 g instead of 1160 g. B divided by the density: $0.150 \times 1000 \div 1.16 \div 40.00 = 3.23$. D dropped the 1000 mL: $0.150 \times 1.16 \div 40.00 = 0.00435$, a thousand times too small.

🧠 **Sense check:** A liter weighs 1160 g, 174 g of it NaOH, and $174 \text{ g} \div 40.00 \text{ g/mol} = 4.35 \text{ mol}$: 4.35 M. ✓

1 · Basis: 1 L soln



2 · Density



3 · Solute



4 · M: ÷ L soln

Worked example 3: molarity to mass percent and molality

Step 1 • Pick a basis

4.50 M H_2SO_4 • density 1.25 g/mL • molar mass 98.08 g/mol

given: molarity and density • wanted: mass percent, then molality • basis: take 1 L of solution

Battery acid is **4.50 M sulfuric acid** with a density of **1.25 g/mL**. What are its mass percent and its molality?

The same liter answers both. Only the denominator changes.

✓ Worked example 3: mass percent

4.50 M H₂SO₄ · density 1.25 g/mL · molar mass 98.08 g/mol

Step 2 · Density links volume and mass

$$1000 \text{ mL soln} \times \frac{1.25 \text{ g soln}}{1 \text{ mL soln}} = 1250 \text{ g solution}$$

Step 3 · Find the solute

Step 4 · Divide by the asked amount

$$\frac{441.4 \text{ g H}_2\text{SO}_4}{1250 \text{ g solution}} \times 100 = 35.3\% \text{ (m/m)}$$

 **Sense check:** The liter weighs 1250 g and carries 441 g of acid, a bit over a third: 35.3%. ✓

1 · Basis: 1 L soln



2 · Density



3 · Solute



4 · %: ÷ g soln

✓ Worked example 3: molality

1 L of 4.50 M H_2SO_4 = 1250 g solution, 441.4 g of it H_2SO_4

A common first attempt

$$m = \frac{4.50 \text{ mol } \text{H}_2\text{SO}_4}{1.250 \text{ kg solution}} = 3.60 \text{ mol/kg} \times$$

Step 4 • Divide by the asked amount

$$m = \frac{4.50 \text{ mol } \text{H}_2\text{SO}_4}{0.809 \text{ kg water}} = 5.56 \text{ mol/kg}$$

 **Sense check:** Under a kilogram of water carries the 4.50 mol, so 5.56 m runs above 4.50 M. ✓

1 • Basis: 1 L soln



2 • Density



3 • Solute



4 • m: ÷ kg solvent



Practice 2

4.00 M NaCl · density 1.15 g/mL · molar mass 58.44 g/mol

basis: take 1 L of solution · wanted: mass percent

A road-deicing brine is **4.00 M sodium chloride** with a density of **1.15 g/mL**. What is its mass percent?

- A. 23.4
- B. 26.9
- C. 234
- D. 20.3

✓ Practice 2 answer: D

4.00 M NaCl · density 1.15 g/mL · molar mass 58.44 g/mol

$$4.00 \text{ mol} \times 58.44 \text{ g/mol} = 233.8 \text{ g NaCl} \cdot 1000 \text{ mL soln} \times \frac{1.15 \text{ g soln}}{1 \text{ mL soln}} = 1150 \text{ g soln}$$

$$\frac{233.8 \text{ g NaCl}}{1150 \text{ g soln}} \times 100 = \mathbf{20.3\% (m/m) \cdot \text{answer D}}$$

A skipped the density: $233.8 \div 1000 \times 100 = 23.4$. B divided by it: $1000 \div 1.15 = 870 \text{ g}$, $233.8 \div 870 \times 100 = 26.9$. C stopped at 234 g of NaCl, a mass, not a percent.

🧠 **Sense check:** The liter weighs 1150 g and carries 234 g of salt, about a fifth: 20.3%. ✓

1 · Basis: 1 L soln



2 · Density



3 · Solute



4 · %: ÷ g soln



Guided example: a solution mixed by mass

18.0 g KNO₃ + 132 g water · density 1.08 g/mL · molar mass 101.11 g/mol
given: the recipe and the density · wanted: molarity · basis: the batch as mixed

A fertilizer solution is made by dissolving **18.0 g of potassium nitrate** in **132 g of water**. The solution's density is **1.08 g/mL**. What is its molarity?

Work all four steps of the method.

✓ Guided example: solution

18.0 g KNO₃ + 132 g water · density 1.08 g/mL · molar mass 101.11 g/mol

Step 1 · Pick a basis

Step 2 · Density links volume and mass

$$18.0 \text{ g} + 132 \text{ g} = 150. \text{ g soln} \times \frac{1 \text{ mL soln}}{1.08 \text{ g soln}} = 138.9 \text{ mL} = \mathbf{0.1389 \text{ L solution}}$$

Step 3 · Find the solute

Step 4 · Divide by the asked amount

$$18.0 \text{ g KNO}_3 \times \frac{1 \text{ mol KNO}_3}{101.11 \text{ g KNO}_3} = 0.178 \text{ mol KNO}_3 \cdot \text{M} = \frac{0.178 \text{ mol KNO}_3}{0.1389 \text{ L solution}} = \mathbf{1.28 \text{ M}}$$

 **Sense check:** Read as 150. mL, the batch gives 1.19 M. Denser than water, it fills only 139 mL: 1.28 M. ✓

1 · Basis: the batch



2 · Density



3 · Solute



4 · M: ÷ L soln



Practice 3

100. g sucrose + 150. mL water · syrup 1.18 g/mL · sucrose 342.30 g/mol
water: 1.00 g/mL · wanted: molarity of the syrup

A canning syrup is made by stirring **100. g of sucrose** into **150. mL of water**. The finished syrup has a density of **1.18 g/mL**. What is its molarity?

- A. 1.95
- B. 0.990
- C. 2.30
- D. 1.17
- E. 1.38

✓ Practice 3 answer: E

100. g sucrose + 150. mL water · syrup 1.18 g/mL · sucrose 342.30 g/mol

$$150. \text{ mL} \times 1.00 \text{ g/mL} = 150. \text{ g water} \cdot 250. \text{ g syrup} \cdot 100. \text{ g} \div 342.30 = 0.2921 \text{ mol}$$

$$250. \text{ g syrup} \times \frac{1 \text{ mL syrup}}{1.18 \text{ g syrup}} = 0.2119 \text{ L} \cdot \text{M} = \frac{0.2921 \text{ mol}}{0.2119 \text{ L}} = \mathbf{1.38 \text{ M} \cdot \text{answer E}}$$

A used the water's 150. mL as the syrup's volume: $0.2921 \div 0.150 = 1.95$. D skipped the density: $0.2921 \div 0.250 = 1.17$. B multiplied by it: $0.2921 \div 0.295 = 0.990$. C left out the sucrose: $0.2921 \div 0.1271 = 2.30$.

🧠 **Sense check:** The sugar adds mass, and the syrup fills 212 mL, more than the 150. mL of water: 1.38 M. ✓

1 · Basis: the batch



2 · Density



3 · Solute



4 · M: ÷ L soln



Practice 4

2.00 M KCl · density 1.09 g/mL · molar mass 74.55 g/mol

basis: take 1 L of solution · wanted: molality

A **2.00 M potassium chloride** solution has a density of **1.09 g/mL**. What is its molality?

- A. 2.13
- B. 2.35
- C. 0.00213
- D. 1.83

✓ Practice 4 answer: A

2.00 M KCl · density 1.09 g/mL · molar mass 74.55 g/mol

basis: 1 L · 2.00 mol KCl = 149.1 g · 1090 g solution · wanted: mol per kg of water

$$1090 \text{ g solution} - 149.1 \text{ g KCl} = 940.9 \text{ g water} = 0.9409 \text{ kg}$$

$$m = \frac{2.00 \text{ mol KCl}}{0.9409 \text{ kg water}} = 2.13 \text{ mol/kg} \cdot \text{answer A}$$

B skipped the density: $1000 - 149.1 = 850.9 \text{ g}$ of water, so 2.35. C kept grams of water: $2.00 \div 940.9 = 0.00213$. D divided by all 1.090 kg of solution: 1.83.

🧠 **Sense check:** 0.941 kg of water, a bit under a kilogram, so 2.13 m sits a bit above 2.00 M. ✓

1 · Basis: 1 L soln



2 · Density



3 · Solute



4 · m: ÷ kg solvent



Practice 5

2.50 m glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) · molar mass 180.16 g/mol

basis: take 1 kg of water · wanted: mass percent

A glucose syrup is **2.50 m** in water. What is its mass percent?

- A. 45.0
- B. 31.1
- C. 450.
- D. 68.9

✓ Practice 5 answer: B

2.50 m glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) · molar mass 180.16 g/mol

$$2.50 \text{ mol} \times \frac{180.16 \text{ g}}{1 \text{ mol}} = 450.4 \text{ g glucose} \cdot 1000 \text{ g water} + 450.4 \text{ g} = 1450.4 \text{ g solution}$$

$$\frac{450.4 \text{ g glucose}}{1450.4 \text{ g solution}} \times 100 = 31.1\% \text{ (m/m)} \cdot \text{answer B}$$

A divided by the water: $450.4 \div 1000 \times 100 = 45.0$. D put the water on top: $1000 \div 1450.4 \times 100 = 68.9$. C stopped at 450. g of glucose, a mass.

🧠 **Sense check:** 1450 g of syrup holds 450 g of glucose, just under a third: 31.1%. ✓

1 · Basis: 1 kg water



2 · no density



3 · Solute



4 · %: ÷ g soln



Practice 6

1.20 m sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) · density 1.12 g/mL · molar mass 342.30 g/mol

basis: take 1 kg of water · wanted: molarity

A sucrose solution for candy making is **1.20 m**, with a density of **1.12 g/mL**. What is its molarity?

- A. 1.34
- B. 0.759
- C. 0.953
- D. 2.28
- E. 0.851

✓ Practice 6 answer: C

1.20 m sucrose ($C_{12}H_{22}O_{11}$) · density 1.12 g/mL · molar mass 342.30 g/mol
 basis: 1000 g water + 1.20 mol sucrose (410.76 g) = 1410.76 g solution · wanted: mol per L

$$1410.76 \text{ g soln} \times \frac{1 \text{ mL soln}}{1.12 \text{ g soln}} = 1.2596 \text{ L} \cdot M = \frac{1.20 \text{ mol}}{1.2596 \text{ L}} = \mathbf{0.953 \text{ M} \cdot \text{answer C}}$$

E skipped the density: $1.20 \div 1.41076 = 0.851$. B multiplied by it: $1410.76 \times 1.12 = 1580$ mL, $1.20 \div 1.580 = 0.759$. A left out the sucrose: $1000 \div 1.12 = 892.9$ mL, $1.20 \div 0.8929 = 1.34$. D subtracted it: $589.24 \text{ g} \div 1.12 = 526.1$ mL, $1.20 \div 0.5261 = 2.28$.

🧠 **Sense check:** The 1.20 mol fill 1.26 L, more than a liter: 0.953 M runs below 1.20 m. ✓

1 · Basis: 1 kg water



2 · Density



3 · Solute



4 · M: ÷ L soln

 Practice 7

48.0% (m/m) HBr · density 1.49 g/mL · molar mass 80.91 g/mol

target: 250. mL of 1.00 M HBr · wanted: mL of the concentrate

A stock bottle reads **48.0% (m/m) HBr, density 1.49 g/mL**. What volume of it, in mL, is diluted with water to make **250. mL of 1.00 M HBr**?

- A. 28.3
- B. 42.1
- C. 2210
- D. 62.8

✓ Practice 7 answer: A

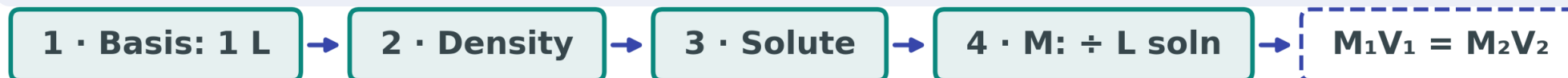
48.0% (m/m) HBr · density 1.49 g/mL · molar mass 80.91 g/mol

$$1000 \text{ mL soln} \times \frac{1.49 \text{ g soln}}{1 \text{ mL soln}} \times \frac{48.0 \text{ g HBr}}{100 \text{ g soln}} \times \frac{1 \text{ mol HBr}}{80.91 \text{ g HBr}} = 8.84 \text{ mol in 1 L: } 8.84 \text{ M}$$

$$V_1 = \frac{M_2 V_2}{M_1} = \frac{1.00 \text{ M} \times 250. \text{ mL}}{8.84 \text{ M}} = 28.3 \text{ mL} \cdot \text{answer A}$$

B skipped the density: 5.93 M, then $250. \div 5.93 = 42.1$. D divided by it: 3.98 M, then $250. \div 3.98 = 62.8$. C flipped the dilution: $8.84 \times 250. \div 1.00 = 2210$.

🧠 **Sense check:** 28.3 mL of 8.84 M acid carries the 0.250 mol of HBr the batch needs. ✓



Check yourself

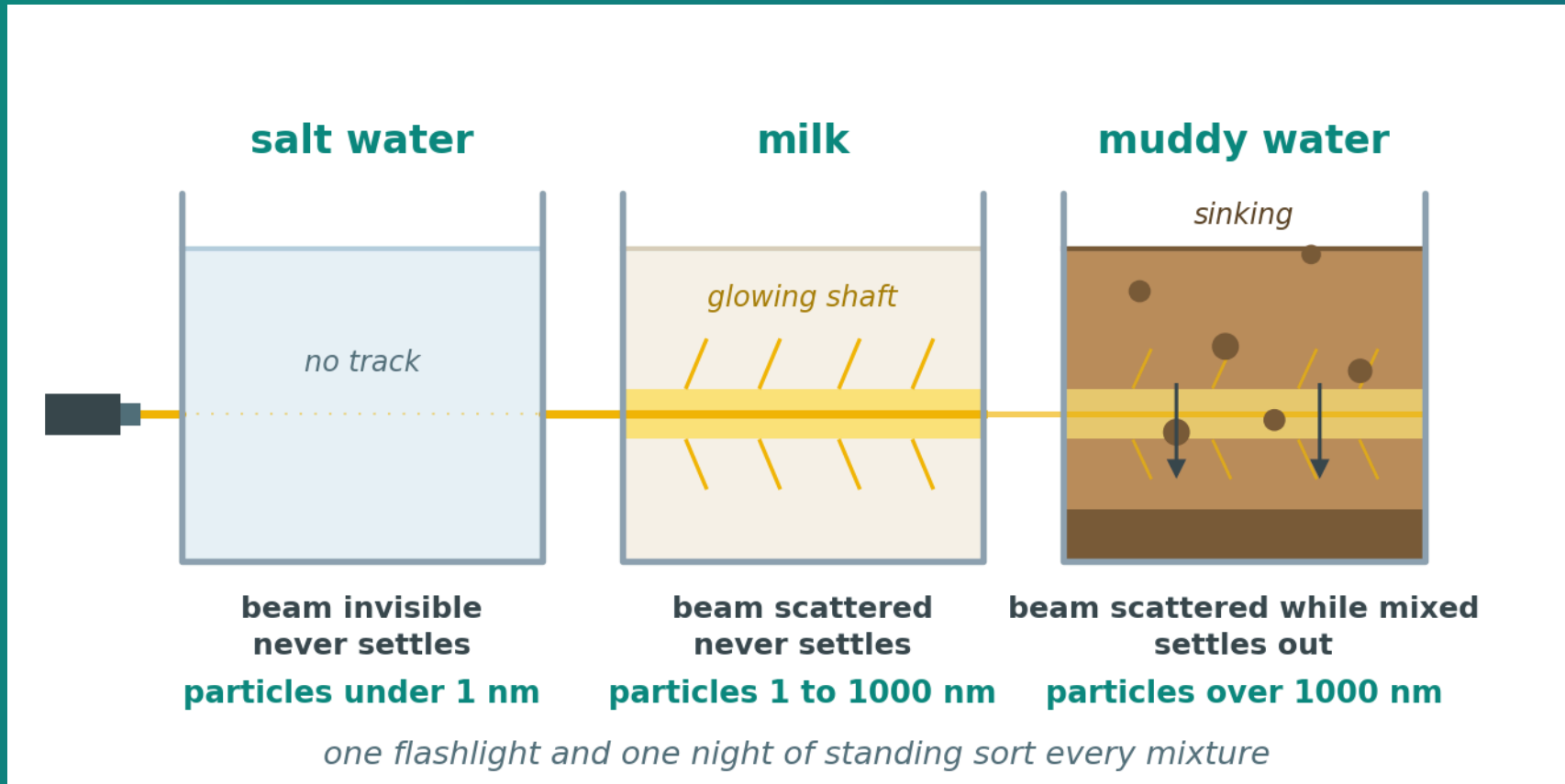
1. A drum reads **36% by mass, density 1.18 g/mL**. Name the two conversion factors that carry that label to a molarity, and say which one makes 1 L weigh more than 1000 g.
2. Molarity and molality both describe one bottle. Which divides the moles by liters of solution, and which by kilograms of solvent, and which of the two needs the density to reach the other?

Freezing-point depression and boiling-point elevation count particles per kilogram of solvent: molality. Turning a labeled molarity or mass percent into molality is the first move in every colligative-property problem.

6 • Colloids

Classify a mixture as a solution, a colloid, or a suspension from particle size, light scattering, and settling, and name a colloid's type from its dispersed phase and its medium.

One flashlight, three mixtures



Fog turns headlight beams into visible shafts; clear air leaves them invisible. Muddy water settles clear overnight; milk stays milk for weeks. Particle size decides both.

Mixtures you already sort

salt water: homogeneous · sand in water: heterogeneous
one uniform phase, even under a microscope · grains visible, settle to the bottom

milk: homogeneous to the eye
under a microscope: fat droplets spread through water

The eye sorts mixtures as homogeneous or heterogeneous. Milk passes for homogeneous, yet its droplets are far larger than molecules. Mixtures like milk form a third class: colloids.

Particle size sets the class

solution: particles under 1 nm

beam invisible · never settles · passes filter paper

colloid: particles 1 to 1000 nm

beam scattered · never settles · passes filter paper

suspension: particles over 1000 nm

cloudy while mixed · settles on standing · caught by filter paper

The size of the dispersed particles decides what a mixture does: whether a beam scatters, whether particles settle, whether filter paper stops them.

memory hook: scatters? not a solution · settles? a suspension

scatters but never settles: a colloid

The Tyndall effect

beam → salt water: crosses unseen

dissolved particles under 1 nm: too small to scatter light

beam → milk: a glowing shaft

colloid particles 1 to 1000 nm: large enough to scatter light in every direction

Colloid particles scatter light sideways, so a beam becomes visible inside the mixture: the **Tyndall effect**. A scattered beam that never settles marks a colloid. A scattered beam that clears on standing marks a suspension.

Colloid types: dispersed phase in medium

fog: liquid in gas → an aerosol · smoke: solid in gas → also an aerosol
whipped cream: gas in liquid → a foam · marshmallow: gas in solid → a solid foam

milk and mayonnaise: liquid in liquid → emulsions · paint: solid in liquid → a sol
gelatin: liquid in a solid protein network → a gel

Naming the dispersed phase and then the medium names the colloid type. Milk scatters light for the same reason fog does: droplets of one phase dispersed in another.

memory hook: "aero" means air, so an aerosol is anything dispersed IN a gas
a foam traps gas inside · an emulsion is liquid in liquid · a sol is solid in liquid

Emulsifying agents

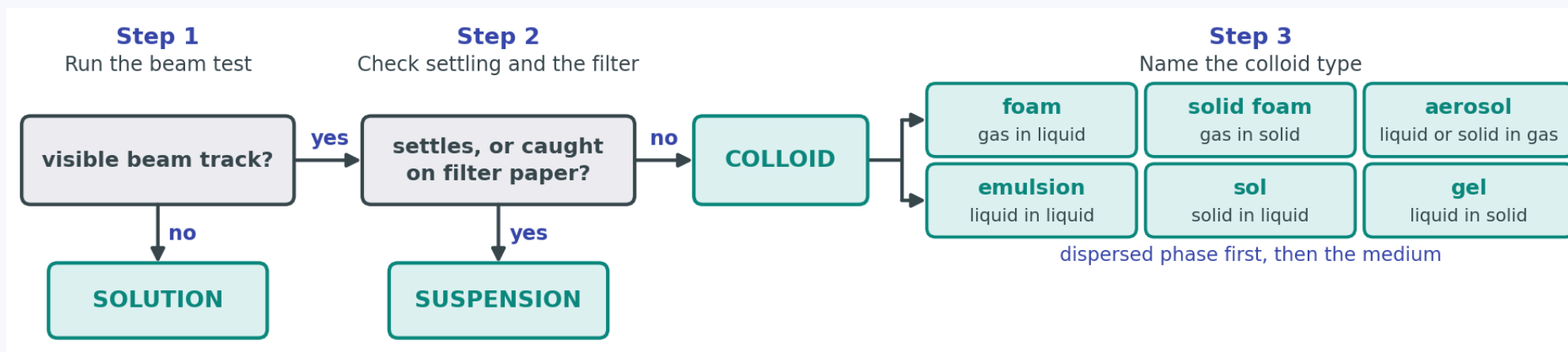
oil + vinegar → two layers · oil + vinegar + egg yolk → mayonnaise
the yolk's lecithin coats every oil droplet: an emulsifying agent

soap: a nonpolar tail on a polar head
tail dissolves into the grease droplet · head faces the water · the droplet stays dispersed

Like dissolves like keeps oil out of vinegar. An **emulsifying agent** sits at the boundary, one end in each liquid, coating droplets so they cannot rejoin. Soap does the same job on grease.

The method

1. **Run the beam test:** no visible track → solution.
2. **Check settling and the filter:** settles or caught on filter paper → suspension. Neither → colloid.
3. **Name the colloid type:** dispersed phase first, then the medium.





Guided example: hand lotion

hand lotion: tiny oil droplets spread through water

laser beam: a bright line · filter paper: nothing caught · a week standing: no settling

Lotion looks perfectly uniform, so a common first call is solution. Classify the lotion. If it is a colloid, name its type.

✓ Guided example: the class

hand lotion: tiny oil droplets spread through water

laser beam: a bright line · filter paper: nothing caught · a week standing: no settling

Step 1 · Run the beam test

Step 2 · Check settling and the filter

bright line · nothing caught · no settling → a colloid

particles between 1 and 1000 nm

 **Sense check:** The lotion scatters light but never settles: a colloid, whatever its uniform look suggests.

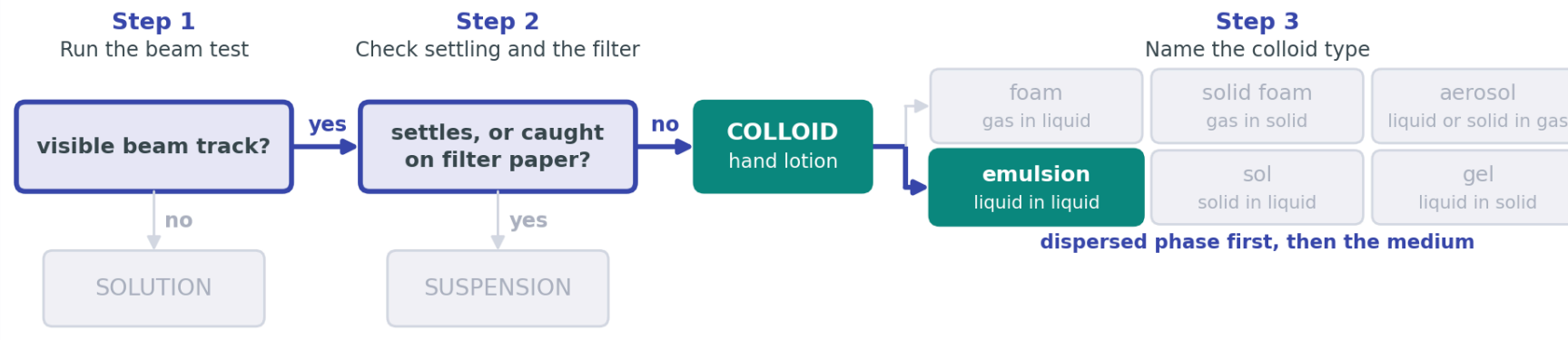
✓ Guided example: the type

hand lotion: tiny oil droplets spread through water
from the tests: bright line · nothing caught · no settling → a colloid

Step 3 · Name the colloid type

oil droplets (liquid) dispersed in water (liquid) → an emulsion
dispersed phase first · then the medium

🧠 **Sense check:** Uniform to the eye, yet a colloid. The beam decides the class; the two phases decide the type.





Practice 1

shaving cream: tiny gas bubbles spread through a liquid soap film
a colloid · wanted: its type

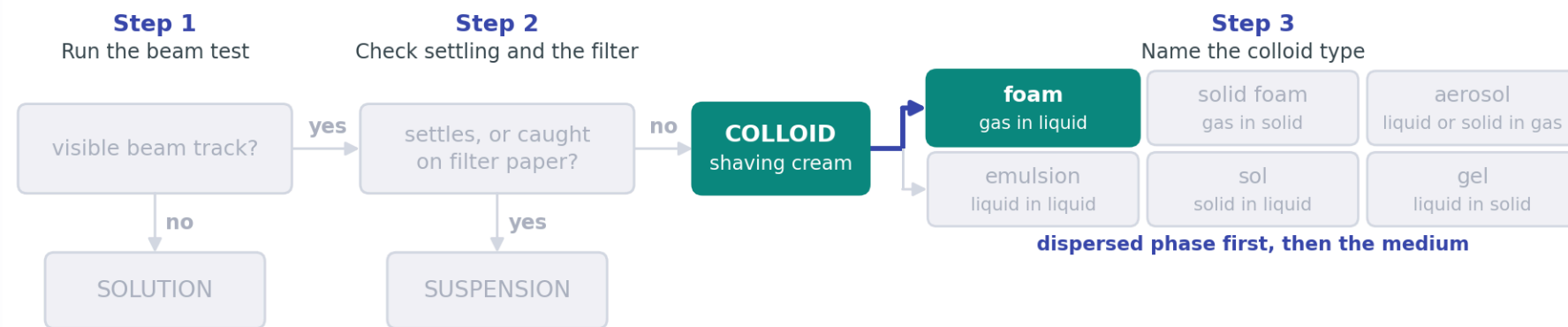
Which colloid type is shaving cream?

- A. A foam
- B. An aerosol
- C. An emulsion
- D. A sol

✓ Practice 1 · answer: A

gas dispersed in a liquid → a foam (answer A)
name the dispersed phase first, then the medium

🧠 **Sense check:** Gas bubbles held in a liquid make a foam, the same type as whipped cream.





Practice 2

sample A: ground coffee stirred into cold water

laser beam: a bright line · filter paper: solid caught · overnight: a layer settles

sample B: egg white stirred into water

laser beam: a bright line · filter paper: nothing caught · a week standing: no settling

Classify each sample from its three test results.

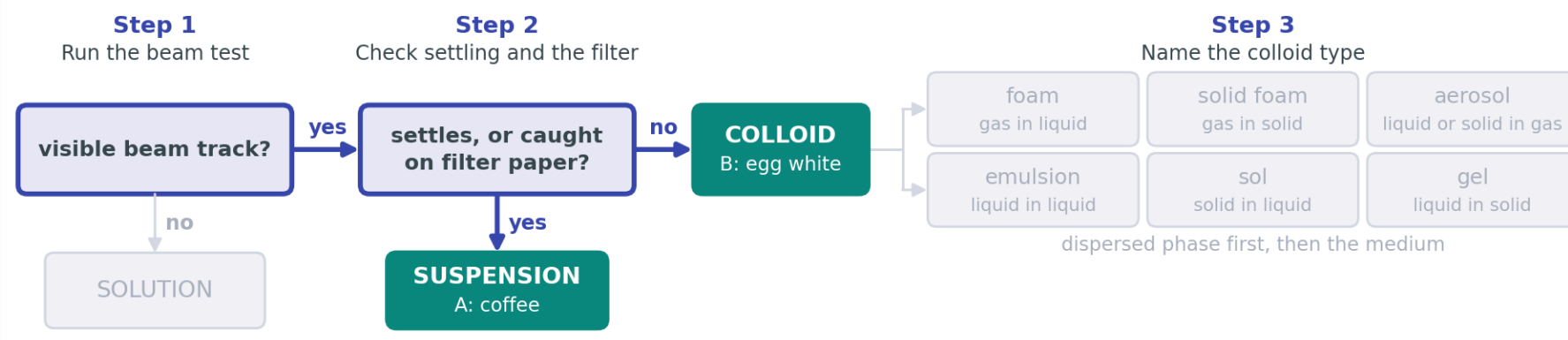
- A. A: colloid · B: suspension
- B. A: suspension · B: colloid
- C. A: colloid · B: colloid
- D. A: suspension · B: solution

✓ Practice 2 · answer: B

A: settles and is caught by filter paper → a suspension · B: scatters but never settles
→ a colloid

answer B · both scatter the beam, so the beam alone cannot tell them apart

🧠 **Sense check:** Run the tests in order for each: the beam rules out a solution for both; settling and the filter paper then split a suspension from a colloid.



Practice 3

A: India ink in water, solid carbon particles spread through it · B: contact-lens saline
laser beam: a glowing shaft in A, unseen in B · filter paper: both pass · a week standing: neither settles

Classify each sample, and name the type of any colloid.

A. A: suspension · B: solution

B. A: colloid, an emulsion · B: solution

C. A: colloid, a sol · B: colloid

D. A: solution · B: solution

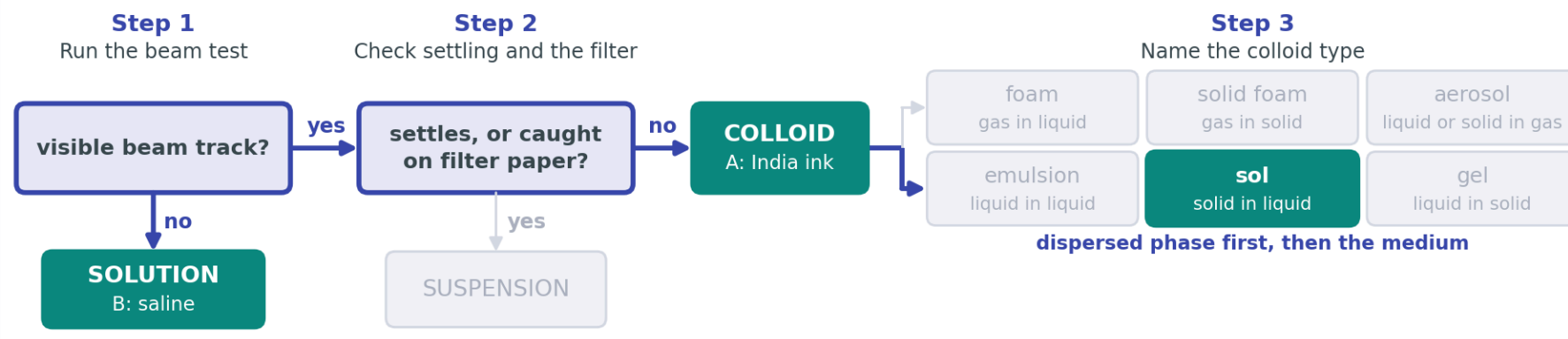
E. A: colloid, a sol · B: solution

✓ Practice 3 · answer: E

A: glowing shaft, passes the filter, never settles → colloid · carbon (solid) in water (liquid) → a sol

B: beam unseen → solution · answer E

🧠 **Sense check:** Beam first for each sample. Only the sample that scatters goes on to a type: solid in liquid, a sol, the same type as paint.





Practice 4

Styrofoam: gas bubbles trapped in solid plastic

a student's call: an aerosol, the same type as smoke, since both pair a solid with a gas

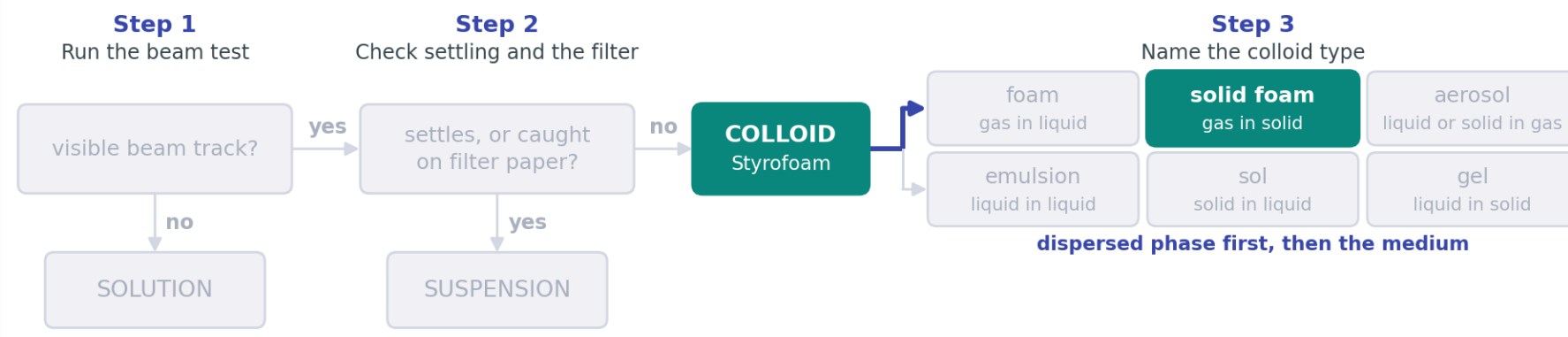
Is the student's call right? If not, which colloid type is Styrofoam?

- A. Right: smoke and Styrofoam are both aerosols
- B. Wrong: gas bubbles make a foam, as in whipped cream
- C. Wrong: a solid that holds its shape is a gel
- D. Wrong: gas dispersed in a solid is a solid foam

✓ Practice 4 · answer: D

Styrofoam: gas dispersed in a solid → a solid foam (answer D)
smoke: solid dispersed in a gas → an aerosol · same two phases, roles reversed

 **Sense check:** Name the dispersed phase first, then the medium. Gas in a solid is a solid foam, the same type as a marshmallow.





Practice 5

salad oil + water, shaken hard
two separate layers again within minutes

Which added substance keeps the oil spread through the water as tiny droplets?

- A. Sodium chloride, NaCl
- B. Sodium stearate, $\text{CH}_3(\text{CH}_2)_{16}\text{COO}^- \text{Na}^+$
- C. Hexane, C_6H_{14}
- D. Glucose, $\text{C}_6\text{H}_{12}\text{O}_6$
- E. NaCl and hexane added together

✓ Practice 5 · answer: B

$\text{CH}_3(\text{CH}_2)_{16}-$ tail: nonpolar, dissolves into the oil droplet · $-\text{COO}^- \text{Na}^+$ head: ionic, faces the water

one end in each liquid → an emulsifying agent (answer B): sodium stearate is a soap

A dissolves only in water, so nothing holds the oil. C dissolves only in the oil. D judged glucose by its carbon chain; its $-\text{OH}$ groups make it polar, so it stays in the water. E split the job between two substances; each one dissolves in its own layer, and no single molecule bridges the boundary.

 **Sense check:** Like dissolves like, one end at a time: a long nonpolar tail on an ionic head coats each droplet.



Practice 6

sample	laser beam	filter paper	standing
diluted white glue in water	bright line	nothing caught	no settling in a week
rubbing alcohol in water	no track	nothing caught	no settling in a week
fine sand in water	bright line	grains caught	a layer within an hour

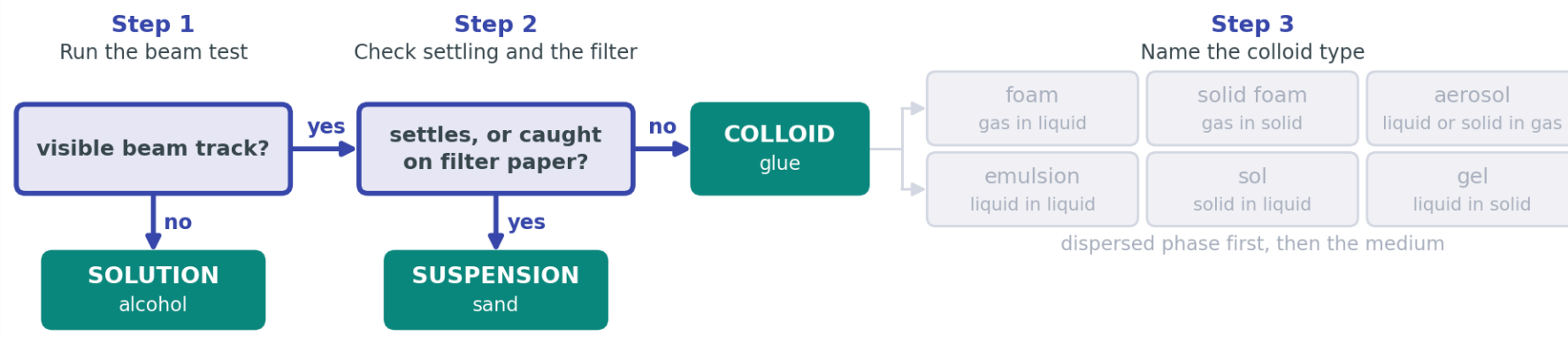
Which line classifies all three mixtures?

- A. glue: solution · alcohol: solution · sand: suspension
- B. glue: suspension · alcohol: solution · sand: suspension
- C. glue: colloid · alcohol: solution · sand: suspension
- D. glue: colloid · alcohol: colloid · sand: suspension
- E. glue: colloid · alcohol: solution · sand: colloid

✓ Practice 6 · answer: C

glue: scatters, passes the filter, never settles → colloid · alcohol: no track → solution
sand: settles and is caught → suspension · answer C · beam first, then settling and the filter paper

🧠 **Sense check:** The beam splits a solution from the rest. Settling and the filter paper then split a suspension from a colloid.



Check yourself

1. A flashlight beam glows inside two different glasses. After a day, one glass has a solid layer on the bottom and the other is unchanged. Classify each mixture.
2. Mayonnaise is mostly oil and vinegar, two liquids that separate on their own. Name the job the egg yolk does, and what each end of such a molecule holds onto.

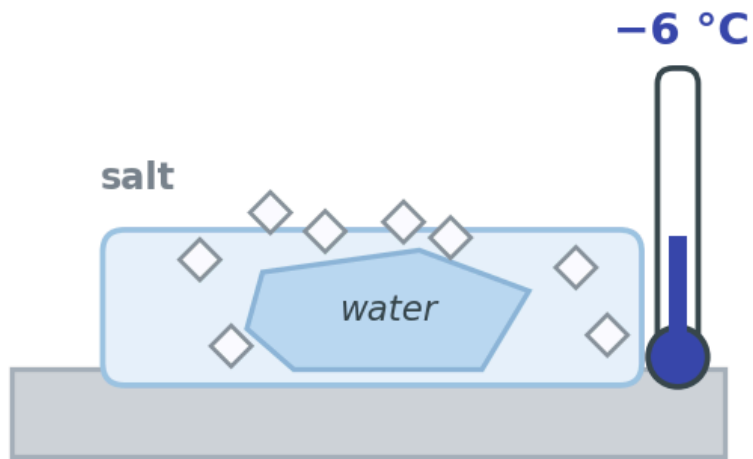
Particle size returns in colligative properties: boiling and freezing points move with the number of dissolved particles, and a colloid's few large particles barely move either one.

7 · 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.

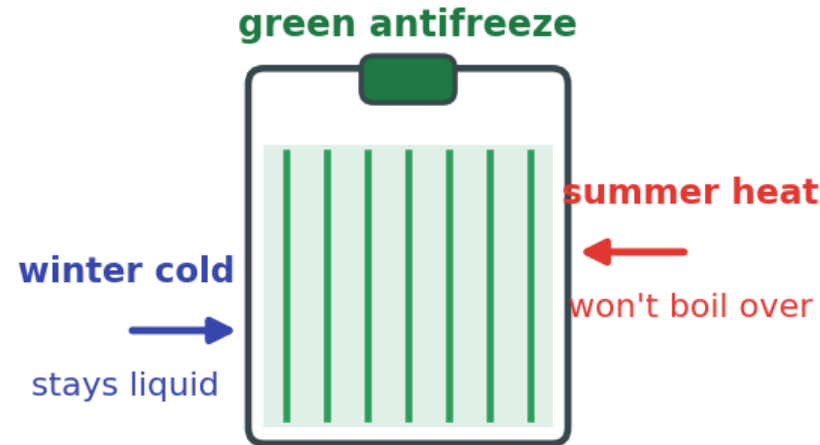
Salt on ice, antifreeze in a radiator

road salt on an icy walk



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

antifreeze in a radiator

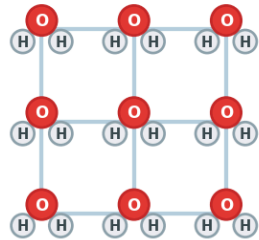


liquid when cold, no boil-over when hot

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.

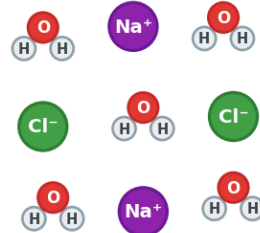
Freezing lower, boiling higher

dissolved particles get in the way of freezing



pure water

molecules lock into an ordered solid



salt water

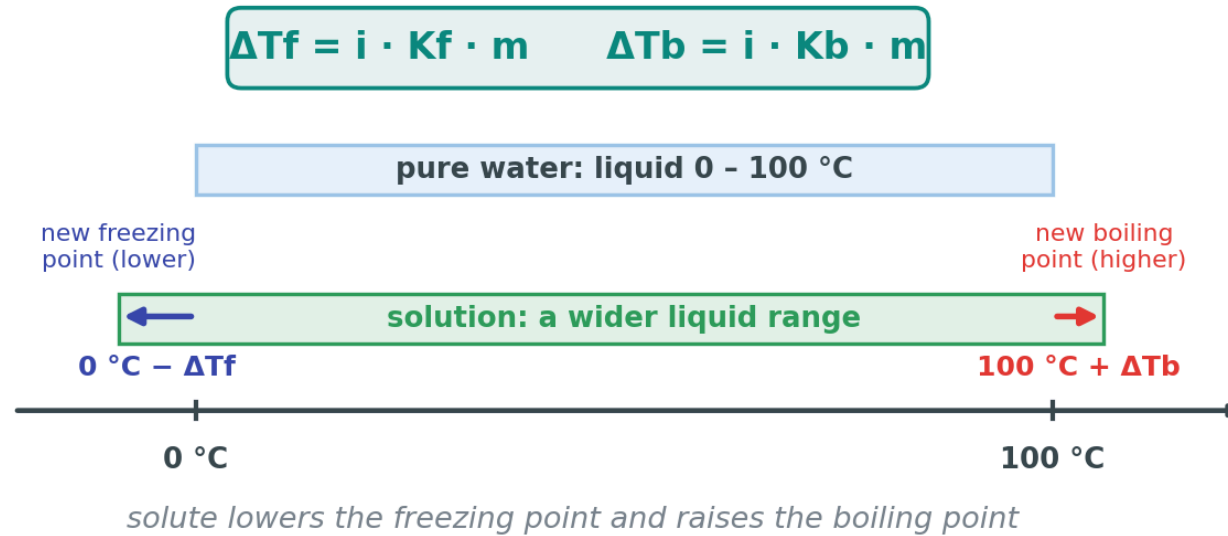
particles block the lattice — must go colder to freeze

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

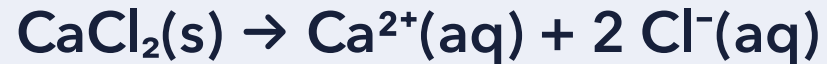
$\Delta T_f = i \cdot K_f \cdot m$ (down) · $\Delta T_b = i \cdot K_b \cdot m$ (up)

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

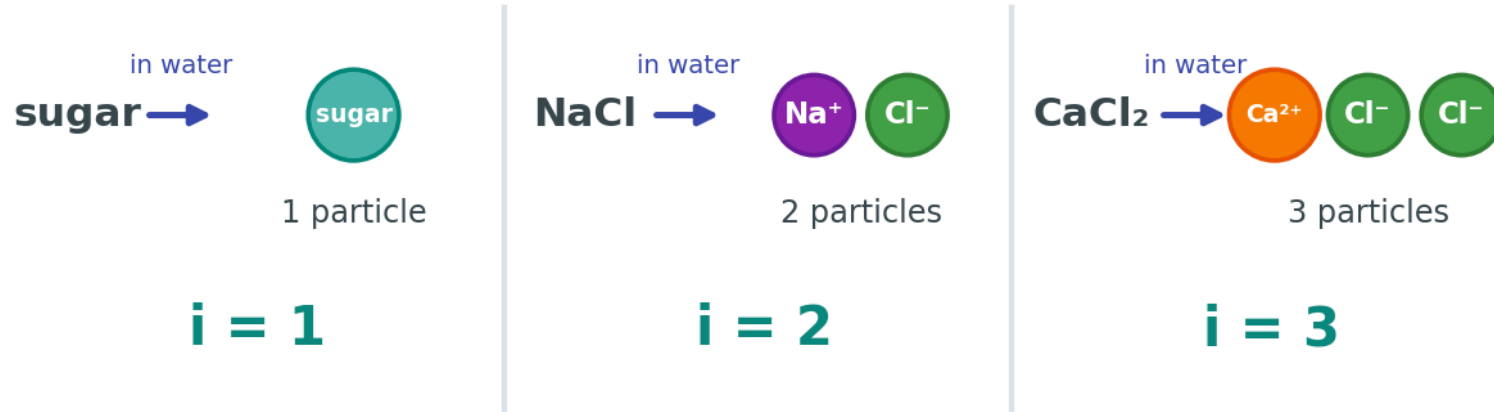


sugar · nonelectrolyte · dissolves whole · 1 particle

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 $100 + \Delta T_b$.



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 $\rightarrow$ mol, $\div$ 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

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

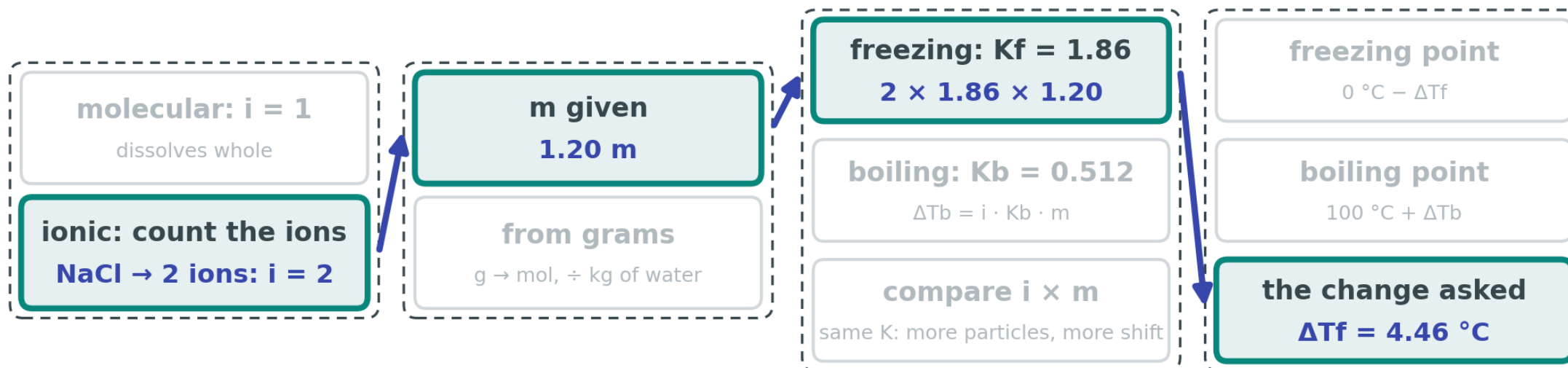
NaCl gives $\text{Na}^+ + \text{Cl}^-$: $i = 2$ · m given · stop at ΔT_f

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

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

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 1 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. ✓



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

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

$i = 3 \cdot \text{molality from grams, } \div 0.500 \text{ kg of water} \cdot \text{the freezing point is asked: shift from } 0 \text{ }^\circ\text{C}$

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.

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

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

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

 **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. ✓

⚠ 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 2

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

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

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 2 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 3

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

$K_b = 0.512\text{ }^{\circ}\text{C}\cdot\text{kg/mol}$ 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 3 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 4

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 4 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. ✓



Practice 5

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

$$\text{molar mass LiCl } 42.39\text{ g/mol} \cdot K_f = 1.86\text{ }^{\circ}\text{C}\cdot\text{kg/mol} \cdot K_b = 0.512\text{ }^{\circ}\text{C}\cdot\text{kg/mol}$$

A desiccant brine is **10.0% lithium chloride by mass** in water. At what temperature, in $^{\circ}\text{C}$, does it freeze?

- A. -8.78
- B. -9.75
- C. 9.75
- D. -4.88
- E. -2.68

✓ Practice 5 answer: B

10.0% by mass: 10.0 g LiCl in every 100 g of solution
 water = 100 g – 10.0 g = 90.0 g · LiCl gives $\text{Li}^+ + \text{Cl}^-$: $i = 2 \cdot K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$

$$10.0 \text{ g} \div 42.39 \text{ g/mol} = 0.2359 \text{ mol LiCl} \rightarrow \frac{0.2359 \text{ mol}}{0.0900 \text{ kg water}} = 2.621 \text{ mol/kg}$$

$$\Delta T_f = 2 \times 1.86 \times 2.621 = 9.75 \text{ }^\circ\text{C} \rightarrow 0 \text{ }^\circ\text{C} - 9.75 \text{ }^\circ\text{C} = \mathbf{-9.75 \text{ }^\circ\text{C} \cdot \text{answer B}}$$

A divided by the whole 100 g of solution: $0.2359 \div 0.100 \text{ kg} = 2.359 \text{ m}$ gives $-8.78 \text{ }^\circ\text{C}$. C stopped at the depression and dropped the sign: $+9.75 \text{ }^\circ\text{C}$ is above $0 \text{ }^\circ\text{C}$. D forgot i : $0 - 1.86 \times 2.621 = -4.88 \text{ }^\circ\text{C}$. E used K_b : $0 - 2 \times 0.512 \times 2.621 = -2.68 \text{ }^\circ\text{C}$.

 **Sense check:** Molality divides by the 90.0 g of water, not the 100 g of solution. Freezing point goes DOWN. ✓



Check yourself

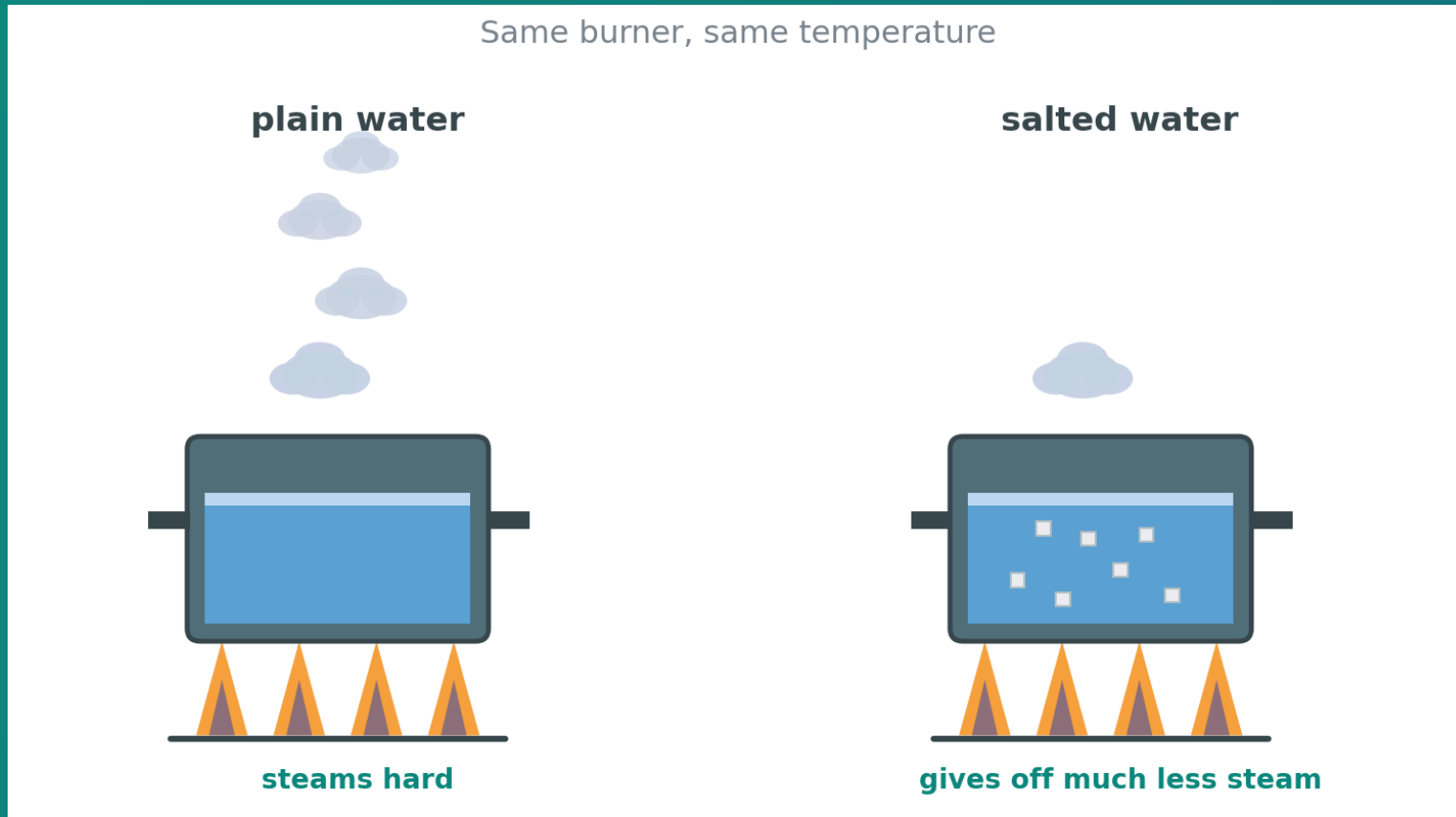
1. A 0.50 molal molecular-solute solution and a 0.50 molal NaCl solution are both cooled. Which freezes at the lower temperature, and why?
2. Write ΔT_b for a 0.750 molal Na₂SO₄ solution, then its boiling point. Which constant belongs in the formula?

Freezing and boiling points depend only on how many particles dissolve. A nonvolatile solute also lowers a solvent's vapor pressure, and Raoult's law sets that pressure from the solvent's mole fraction.

8 • Raoult's Law

Use Raoult's law to find a solution's vapor pressure from the moles of solvent and nonvolatile solute, scaling the pure solvent's pressure by the solvent's mole fraction, and confirm the result falls below the pure value.

Salted water, less steam

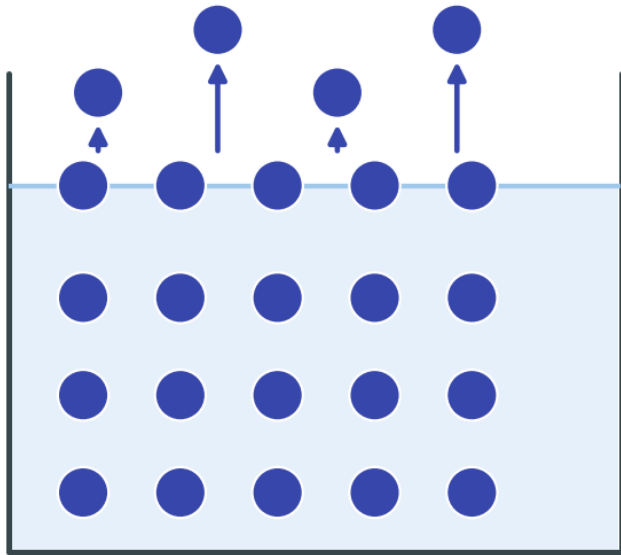


Two pots reach the same temperature on one burner. The pot of plain water steams hard; the salted pot gives off noticeably less. The dissolved salt holds the water back.

A nonvolatile solute lowers vapor pressure

A nonvolatile solute lowers the vapor pressure

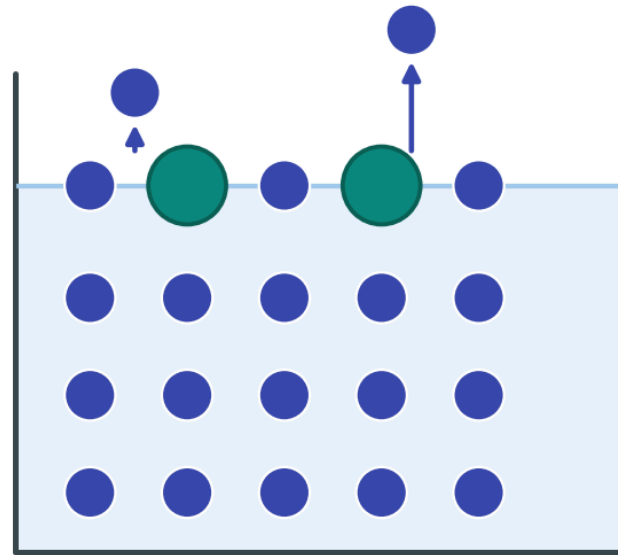
pure solvent



every surface spot is solvent → many escape

solution: solvent + nonvolatile solute

● solvent ● nonvolatile solute



solute holds surface spots → fewer escape

Vapor comes only from solvent at the surface. A nonvolatile solute takes up surface spots and never leaves, so the pressure drops with the solvent's fraction.

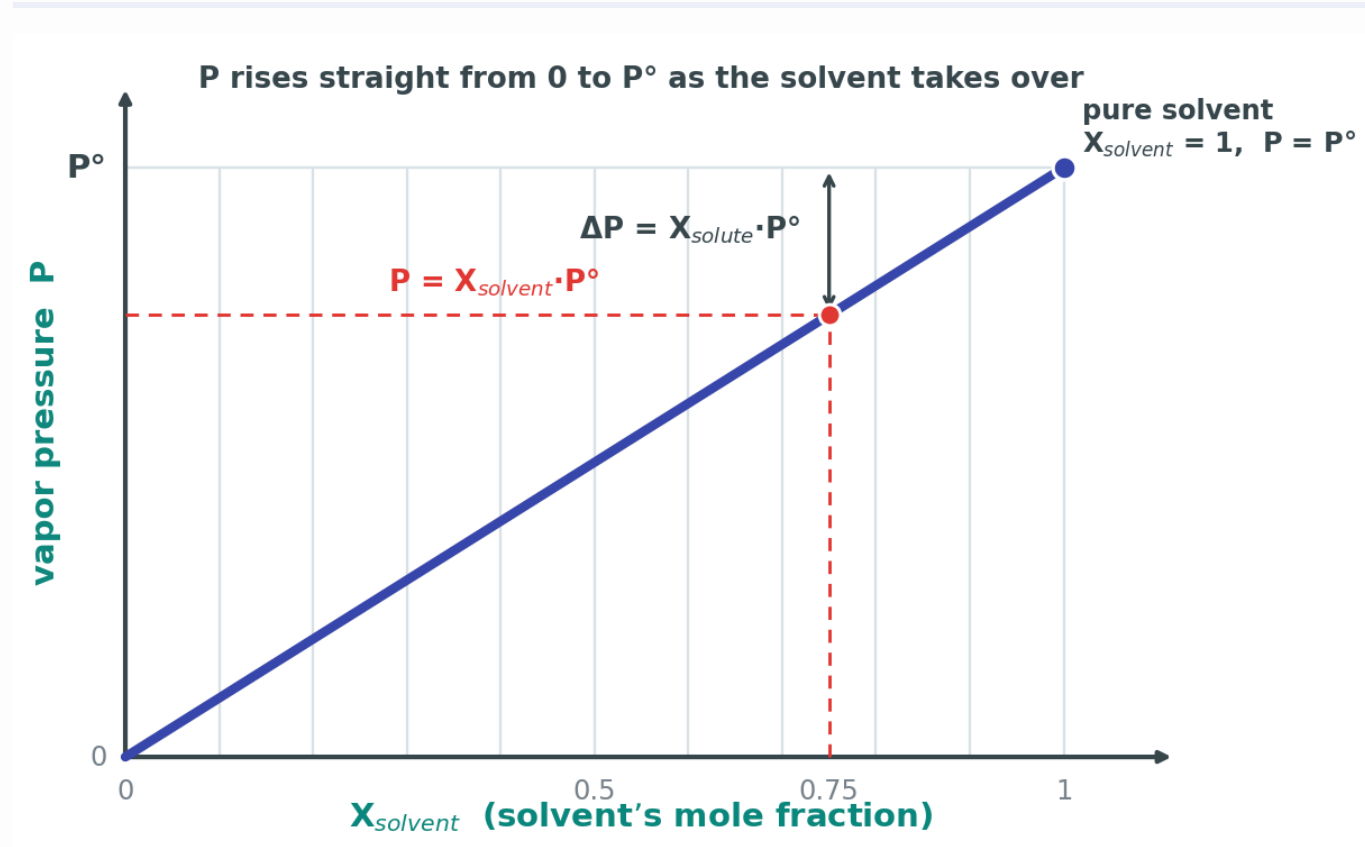
Mole fraction scales a pressure

Dalton's law: $P_{\text{gas}} = \text{mole fraction} \times P_{\text{total}}$

$4.0 \text{ mol N}_2 + 1.0 \text{ mol O}_2 \cdot \text{mole fraction of O}_2 = 1.0 \div 5.0 = 0.20 \cdot P(\text{O}_2) = 0.20 \times 3.0 \text{ atm} = 0.60 \text{ atm}$

In a gas mixture, each gas's mole fraction sets its share of the total pressure. A solution works the same way. The solvent's mole fraction sets its share of P° , the pure liquid's vapor pressure.

Raoult's law



The solution's vapor pressure is the pure value scaled by the solvent's mole fraction: $P = X_{\text{solvent}} \cdot P^\circ$. Pure solvent sits at P° ; every added mole of solute pulls P straight down the line.

The pressure lowering, ΔP

$$\Delta P = X_{\text{solute}} \cdot P^\circ = P^\circ - P$$

$$X_{\text{solute}} = \text{mol solute} \div \text{total moles} \cdot X_{\text{solvent}} + X_{\text{solute}} = 1$$

The drop from the pure pressure to the solution's is the pressure lowering, ΔP . It equals the solute's mole fraction times P° , the same amount you get as P° minus the solution's pressure.

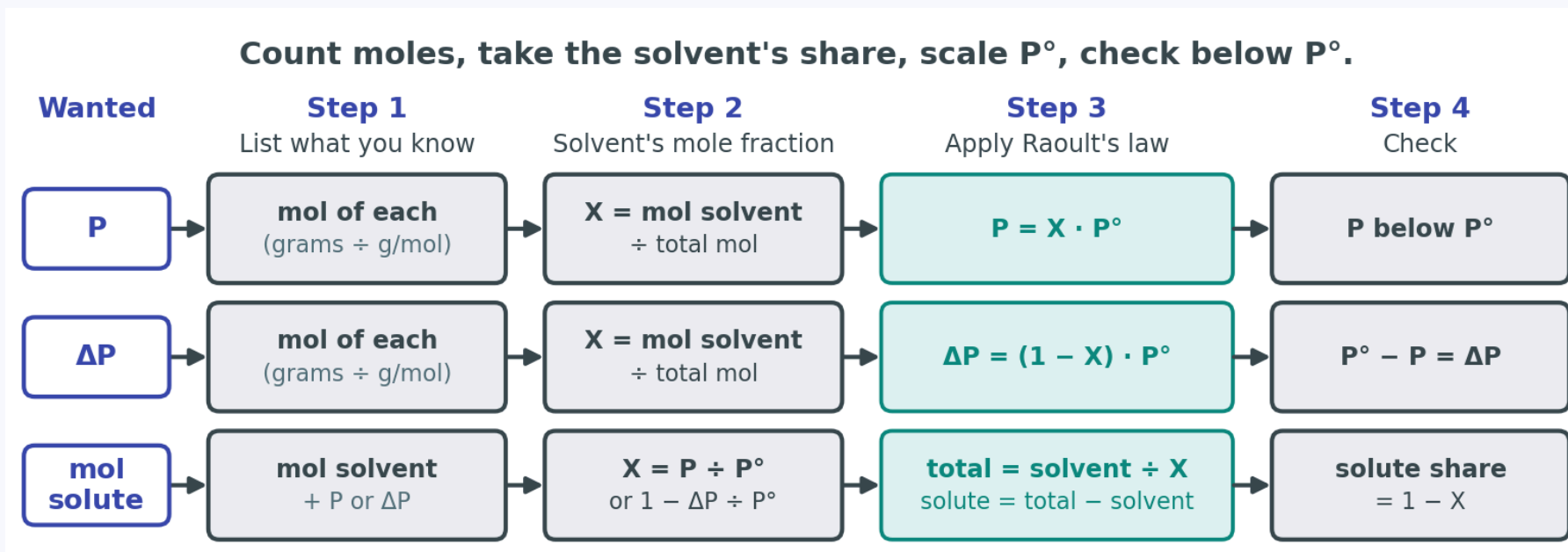
memory hook: P follows the solvent, ΔP follows the solute

$$P = X_{\text{solvent}} \cdot P^\circ \cdot \Delta P = X_{\text{solute}} \cdot P^\circ \cdot \text{together they rebuild } P^\circ$$



The method

1. **List what you know:** moles of each and P° . Mark the unknown.
2. **Solvent's mole fraction:** $X_{\text{solvent}} = \text{mol solvent} \div \text{total}$.
3. **Apply Raoult's law:** $P = X_{\text{solvent}} \times P^\circ$; solve for the unknown.
4. **Check:** P below P° .





Worked example 1: vapor pressure of a solution

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

given: 5.00 mol water · 1.00 mol nonvolatile solute · $P^{\circ} = 23.8$ torr · wanted: P

Sugar is a nonvolatile solute: it dissolves in water but never enters the vapor. Dissolve 1.00 mol of it in 5.00 mol of water. Find the solution's vapor pressure. (*pure water: 23.8 torr*)

List what you know, then take the solvent's mole fraction.

✓ Worked example 1: solution

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

given: 5.00 mol water · 1.00 mol nonvolatile solute · $P^{\circ} = 23.8$ torr · wanted: P

Step 1 · List what you know

Step 2 · Solvent's mole fraction

$$X_{\text{solvent}} = \frac{5.00 \text{ mol water}}{5.00 + 1.00 \text{ mol total}} = 0.833$$

Step 3 · Apply Raoult's law

$$P = 0.833 \times 23.8 \text{ torr} = 19.8 \text{ torr}$$

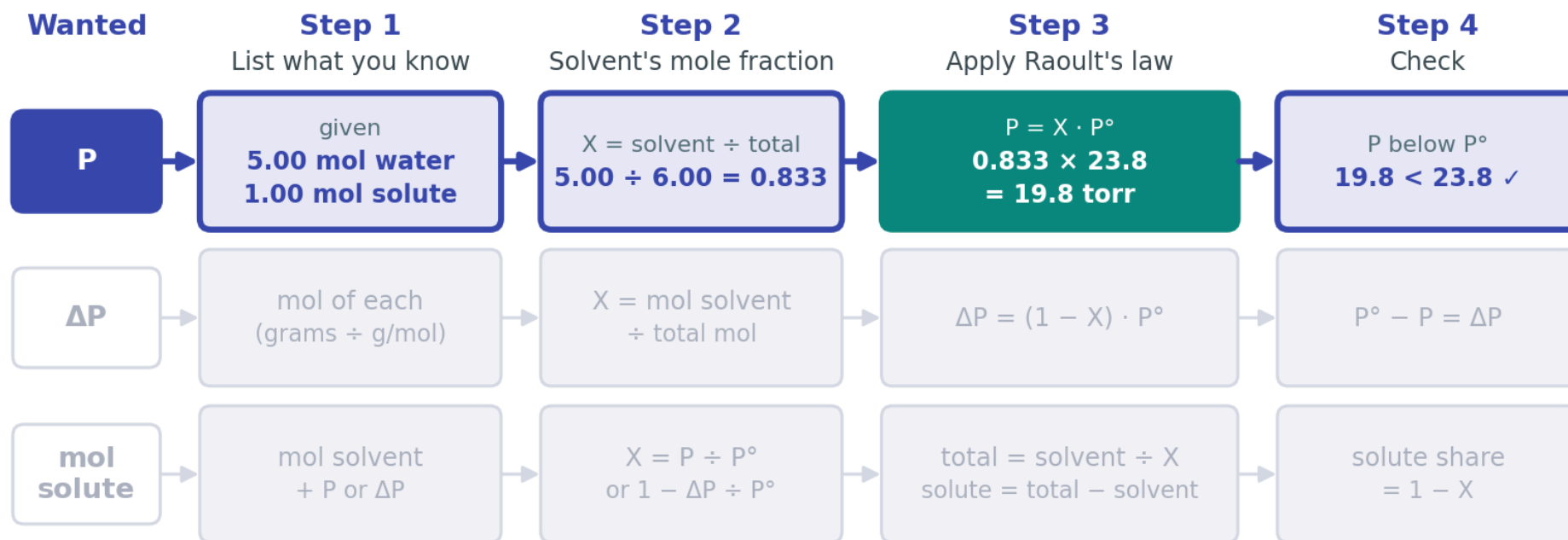
Step 4 · Check

🧠 **Sense check:** 19.8 torr sits below the pure 23.8 torr: the nonvolatile solute lowered it. ✓

✓ Worked example 1: the route on the map

$$P = X_{\text{solvent}} \cdot P^{\circ} = 0.833 \times 23.8 \text{ torr} = 19.8 \text{ torr}$$

given: 5.00 mol water · 1.00 mol nonvolatile solute · $P^{\circ} = 23.8 \text{ torr}$ · found: $P = 19.8 \text{ torr}$



🧠 **Sense check:** The pressure is wanted, so the top lane: the solvent's share of the moles, times P° . ✓



Worked example 2: the pressure lowering

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

given: 6.00 mol benzene · 2.00 mol nonvolatile solute · $P^{\circ} = 95.1$ torr · wanted: ΔP and P

A nonvolatile solute is dissolved in benzene: 2.00 mol of solute in 6.00 mol of benzene. Find the pressure lowering, ΔP , and the solution's vapor pressure. (*pure benzene: 95.1 torr*)

List the moles, then take each mole fraction.

✓ Worked example 2: solution

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

given: 6.00 mol benzene · 2.00 mol nonvolatile solute · $P^{\circ} = 95.1$ torr · wanted: ΔP and P

Step 1 · List what you know

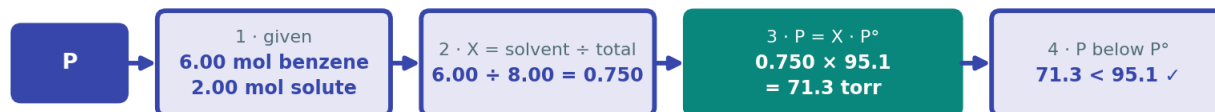
Step 2 · Solvent's mole fraction

$$X_{\text{solvent}} = \frac{6.00 \text{ mol benzene}}{6.00 + 2.00 \text{ mol total}} = 0.750$$

Step 3 · Apply Raoult's law

$$P = 0.750 \times 95.1 \text{ torr} = 71.3 \text{ torr}$$

🧠 **Sense check:** 71.3 torr is below the pure 95.1 torr. ✓



✓ Worked example 2: the pressure lowering

$$\Delta P = X_{\text{solute}} \cdot P^{\circ} = P^{\circ} - P$$

$$2.00 \text{ mol solute of } 8.00 \text{ mol total} \cdot X_{\text{solute}} = 0.250 \cdot P^{\circ} = 95.1 \text{ torr} \cdot P = 71.3 \text{ torr}$$

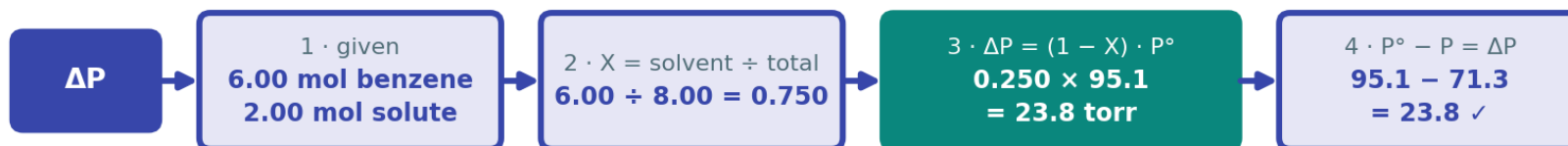
The lowering ΔP

$$\Delta P = X_{\text{solute}} \cdot P^{\circ} = 0.250 \times 95.1 = \mathbf{23.8 \text{ torr}}$$

Step 4 • Check

$$P^{\circ} - P = 95.1 - 71.3 = \mathbf{23.8 \text{ torr}}$$

🧠 **Sense check:** Both routes to ΔP agree, and P stays below P° . ✓



Your turn: an ethanol solution

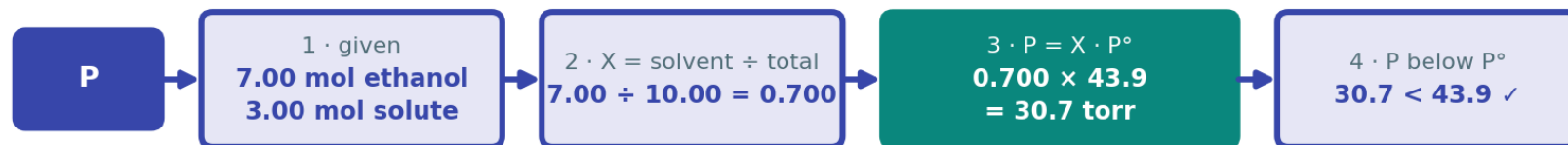
$$P = X_{\text{solvent}} \cdot P^{\circ}$$

given: 7.00 mol ethanol · 3.00 mol nonvolatile solute · $P^{\circ} = 43.9$ torr · wanted: P

$$X_{\text{solvent}} = \frac{\text{mol ethanol}}{\text{mol total}} = \text{ } \rightarrow P = \text{ } \times 43.9 = \text{ } \text{ torr}$$

$$X_{\text{solvent}} = \frac{7.00 \text{ mol ethanol}}{10.00 \text{ mol total}} = 0.700 \rightarrow P = 0.700 \times 43.9 = \mathbf{30.7 \text{ torr}}$$

 **Sense check:** 30.7 torr is below the pure 43.9 torr. The solute takes the rest: $43.9 - 30.7 = 13.2$ torr of lowering. ✓



⚠ Where this goes wrong

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

$$5.00 \text{ mol water} \cdot 1.00 \text{ mol solute} \cdot P^{\circ} = 23.8 \text{ torr} \cdot \text{correct } P = 19.8 \text{ torr}$$

Using the solute's mole fraction. $(1.00 \div 6.00) \times 23.8 = 3.97$ torr is the pressure LOWERING, ΔP : the amount removed. The pressure that remains uses the solvent's fraction: $(5.00 \div 6.00) \times 23.8 = 19.8$ torr.

Assuming the solute changes nothing. Leaving the answer at 23.8 torr. A nonvolatile solute always lowers the vapor pressure, so the result must fall below 23.8 torr, here to 19.8 torr.

Skipping the total in the fraction. $(1.00 \div 5.00) \times 23.8 = 4.76$ torr divides solute by solvent. A mole fraction divides by the TOTAL moles: the solvent's fraction is $5.00 \div 6.00$, not $1.00 \div 5.00$.



Guided example: vapor pressure from grams

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

given: 46.0 g glycerol (92.09 g/mol) · 138 g ethanol (46.07 g/mol) · $P^{\circ} = 43.9$ torr · wanted: P

A hand-sanitizer base mixes 46.0 g of glycerol, a nonvolatile liquid, into 138 g of ethanol. Find the solution's vapor pressure. (*pure ethanol: 43.9 torr*)

Convert both masses to moles, then take the solvent's mole fraction.

✓ Guided example: solution

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

given: 46.0 g glycerol (92.09 g/mol) · 138 g ethanol (46.07 g/mol) · $P^{\circ} = 43.9$ torr · wanted: P

Step 1 · List what you know

$$46.0 \text{ g glycerol} \times \frac{1 \text{ mol}}{92.09 \text{ g}} = 0.500 \text{ mol glycerol}$$

$$138 \text{ g ethanol} \times \frac{1 \text{ mol}}{46.07 \text{ g}} = 3.00 \text{ mol ethanol}$$

🧠 **Sense check:** Glycerol has a third of the mass but a sixth of the moles: $3.00 \div 0.500 = 6.00$. ✓

✓ Guided example: the vapor pressure

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

$$3.00 \text{ mol ethanol} \cdot 0.500 \text{ mol glycerol} \cdot P^{\circ} = 43.9 \text{ torr} \cdot \text{wanted: } P$$

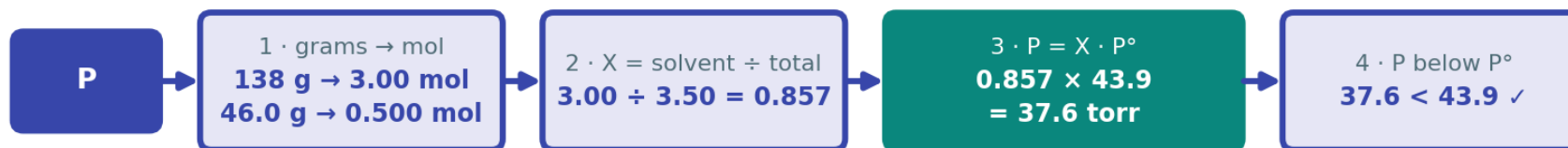
Step 2 • Solvent's mole fraction

Step 3 • Apply Raoult's law

$$X_{\text{solvent}} = \frac{3.00 \text{ mol ethanol}}{3.00 + 0.500 \text{ mol total}} = 0.857 \rightarrow P = 0.857 \times 43.9 \text{ torr} = \mathbf{37.6 \text{ torr}}$$

Step 4 • Check

🧠 Sense check: 37.6 torr sits below the pure 43.9 torr. ✓



 Practice 1

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

given: 68.4 g sucrose (342.30 g/mol) · 90.0 g water (18.02 g/mol) · $P^{\circ} = 23.8$ torr · wanted: P

A syrup is made from **68.4 g of sucrose** (342.30 g/mol) and **90.0 g of water** (18.02 g/mol). What is its vapor pressure, in torr? (*pure water at 25 °C: 23.8 torr*)

- A. 13.5
- B. 0.916
- C. 1.62
- D. 22.9

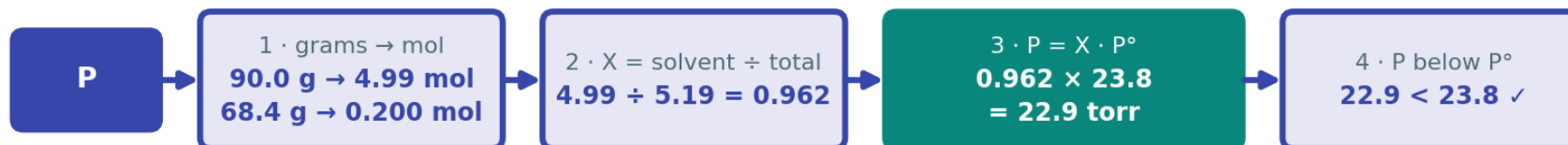
✓ Practice 1 answer: D

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

$$68.4 \text{ g sucrose} \rightarrow 0.200 \text{ mol} \cdot 90.0 \text{ g water} \rightarrow 4.99 \text{ mol} \cdot \text{total} = 5.19 \text{ mol} \cdot P^{\circ} = 23.8 \text{ torr}$$

$$X_{\text{solvent}} = \frac{4.99 \text{ mol water}}{(4.99 + 0.200) \text{ mol}} = 0.962 \rightarrow P = 0.962 \times 23.8 = \mathbf{22.9 \text{ torr} \cdot \text{answer D}}$$

 **Sense check:** Sucrose is heavy, so 68.4 g is only 0.200 mol against 4.99 mol of water: a small lowering, 22.9 torr just under 23.8. ✓





Worked example 3: back out the moles of solute

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

given: 8.0 mol benzene · solution $P = 76.08$ torr · $P^{\circ} = 95.1$ torr · wanted: mol solute

A nonvolatile solute in 8.0 mol of benzene gives a solution vapor pressure of 76.08 torr. How many moles of solute are dissolved? (*pure benzene: 95.1 torr*)

A common first attempt: read the pressure ratio $76.08 \div 95.1$ as the solute's share. Test it.

✓ Worked example 3: solution

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

given: 8.0 mol benzene · solution $P = 76.08$ torr · $P^{\circ} = 95.1$ torr · wanted: mol solute

A common first attempt

$$X_{\text{solute}} ? = \frac{76.08 \text{ torr}}{95.1 \text{ torr}} = 0.800 \rightarrow \text{solute} = \mathbf{32 \text{ mol} \times}$$

Step 1 · List what you know

Step 2 · Solvent's mole fraction

$$X_{\text{solvent}} = \frac{P}{P^{\circ}} = \frac{76.08 \text{ torr}}{95.1 \text{ torr}} = \mathbf{0.800}$$

 **Sense check:** The pressure ratio is the solvent's share; its complement, 0.200, belongs to the solute. ✓

✓ Worked example 3: moles of solute

$$X_{\text{solvent}} = \text{mol benzene} \div \text{total}$$

$$8.0 \text{ mol benzene} \cdot X_{\text{solvent}} = 0.800 \cdot P = 76.08 \text{ torr} \cdot P^{\circ} = 95.1 \text{ torr}$$

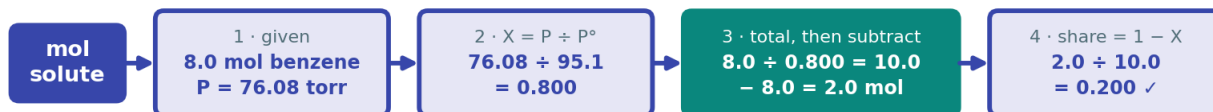
Step 3 • Apply Raoult's law

$$\text{total} = \frac{8.0 \text{ mol benzene}}{0.800} = 10.0 \text{ mol} \rightarrow \text{solute} = 10.0 - 8.0 = \mathbf{2.0 \text{ mol}}$$

Step 4 • Check

$$\text{solute share} = 1 - 0.800 = 0.200 \rightarrow 0.200 \times 10.0 = \mathbf{2.0 \text{ mol} \checkmark}$$

 **Sense check:** Both fractions give 2.0 mol; the 32 mol read the solvent's share as the solute's. ✓



Take-home: use the solvent's mole fraction

solute's fraction → the lowering: $(1.00 \div 6.00) \times 23.8 = 3.97 \text{ torr} \times$
 $5.00 \text{ mol water} \cdot 1.00 \text{ mol solute} \cdot P^\circ = 23.8 \text{ torr} \cdot$ this is ΔP , not the pressure

solvent's fraction → the pressure: $(5.00 \div 6.00) \times 23.8 = 19.8 \text{ torr} \checkmark$
 $P = X_{\text{solvent}} \cdot P^\circ \cdot$ the two add back to P° : $3.97 + 19.8 = 23.8 \text{ torr}$

Raoult's law multiplies P° by the SOLVENT's mole fraction. The solute's fraction gives the drop, ΔP : the amount removed, not the pressure that remains.

 Practice 2

$$\Delta P = X_{\text{solute}} \cdot P^{\circ}$$

given: 12.0 mol acetone · lowering $\Delta P = 46.2$ torr · $P^{\circ} = 231$ torr · wanted: mol solute

A nonvolatile solute dissolved in 12.0 mol of acetone lowers the vapor pressure by 46.2 torr. How many moles of solute are dissolved? (*pure acetone: 231 torr*)

- A. 3.00
- B. 15.0
- C. 48.0
- D. 60.0

✓ Practice 2 answer: A

$$\Delta P = X_{\text{solute}} \cdot P^{\circ}$$

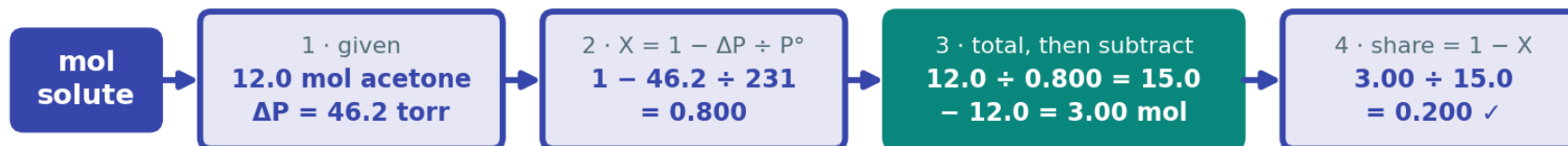
$$12.0 \text{ mol acetone} \cdot \Delta P = 46.2 \text{ torr} \cdot P^{\circ} = 231 \text{ torr}$$

$$X_{\text{solute}} = \frac{46.2 \text{ torr}}{231 \text{ torr}} = 0.200 \rightarrow X_{\text{solvent}} = 0.800 \rightarrow \text{total} = \frac{12.0 \text{ mol}}{0.800} = 15.0 \text{ mol}$$

$$\rightarrow \text{solute} = 15.0 - 12.0 = \mathbf{3.00 \text{ mol} \cdot \text{answer A}}$$

🧠 **Sense check:** 3.00 mol of solute in 15.0 mol total is a 0.200 share, and $0.200 \times 231 = 46.2 \text{ torr}$, the lowering given.

✓





Practice 3

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

given: 215 g chloroform (CHCl_3 , 119.38 g/mol) · solution $P = 157.6$ torr · $P^{\circ} = 197.0$ torr · wanted: mol solute

A nonvolatile solute is dissolved in **215 g of chloroform** (CHCl_3 , 119.38 g/mol). The solution's vapor pressure is **157.6 torr**. How many moles of solute are dissolved? (*pure chloroform: 197.0 torr*)

- A. 2.25
- B. 53.8
- C. 0.360
- D. 0.450
- E. 7.20

✓ Practice 3 answer: D

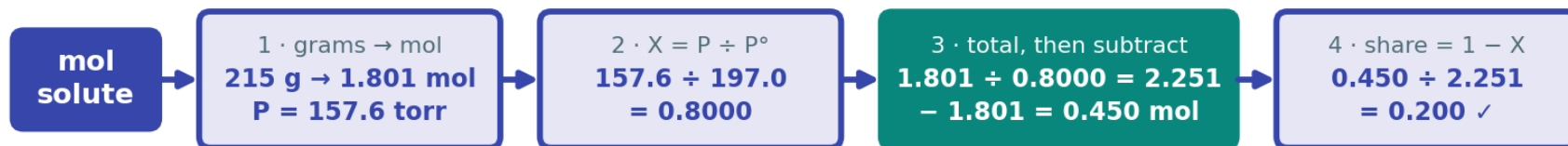
$$P = X_{\text{solvent}} \cdot P^{\circ}$$

215 g chloroform (119.38 g/mol) $\rightarrow$ 1.801 mol $\cdot$ $P = 157.6$ torr $\cdot$ $P^{\circ} = 197.0$ torr $\cdot$ wanted: mol solute

$$X_{\text{solvent}} = \frac{157.6 \text{ torr}}{197.0 \text{ torr}} = 0.8000 \rightarrow \text{total} = \frac{1.801 \text{ mol CHCl}_3}{0.8000} = 2.251 \text{ mol}$$

$$\text{solute} = 2.251 \text{ mol} - 1.801 \text{ mol CHCl}_3 = \mathbf{0.450 \text{ mol} \cdot \text{answer D}}$$

🧠 **Sense check:** 0.450 of 2.251 mol is a 0.200 share: $0.200 \times 197.0 = 39.4$ torr, the drop to 157.6 torr. ✓





Practice 4

$$\Delta P = X_{\text{solute}} \cdot P^{\circ}$$

given: 31.0 g ethylene glycol (62.07 g/mol) · 81.0 g water (18.02 g/mol) · $P^{\circ} = 31.8$ torr · wanted: ΔP

A coolant blend holds **31.0 g of ethylene glycol** (62.07 g/mol) in **81.0 g of water** (18.02 g/mol). By how much, in torr, does its vapor pressure fall below pure water's? (*pure water at 30 °C: 31.8 torr*)

- A. 28.6
- B. 8.80
- C. 3.18
- D. 3.53

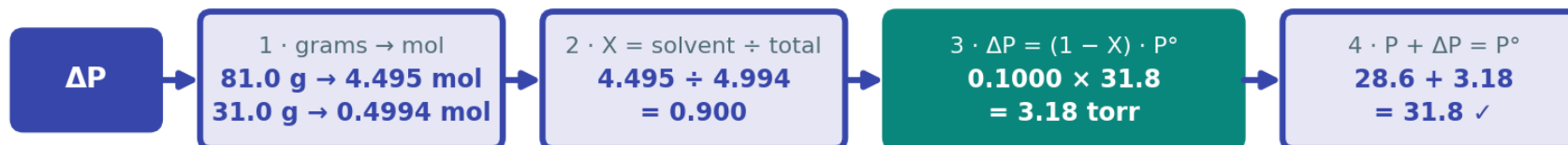
✓ Practice 4 answer: C

$$\Delta P = X_{\text{solute}} \cdot P^{\circ}$$

31.0 g ethylene glycol $\rightarrow$ 0.4994 mol $\cdot$ 81.0 g water $\rightarrow$ 4.495 mol $\cdot$ total = 4.994 mol $\cdot$ $P^{\circ} = 31.8$ torr

$$X_{\text{solute}} = \frac{0.4994 \text{ mol glycol}}{(0.4994 + 4.495) \text{ mol}} = 0.1000 \rightarrow \Delta P = 0.1000 \times 31.8 = \mathbf{3.18 \text{ torr} \cdot \text{answer C}}$$

 **Sense check:** The glycol is a tenth of the moles, so the drop is a tenth of 31.8 torr: 3.18 torr. The 28.6 torr left is the solution's pressure. ✓



 Practice 5

$$P = X_{\text{solvent}} \cdot P^{\circ}$$

given: 32.5% urea by mass (60.06 g/mol) in water (18.02 g/mol) · $P^{\circ} = 17.5$ torr · wanted: P

A tank of diesel exhaust fluid is **32.5% urea by mass** (60.06 g/mol) in water. Find its vapor pressure at 20 °C, in torr. (*pure water at 20 °C: 17.5 torr*)

- A. 15.9
- B. 15.3
- C. 2.21
- D. 11.8

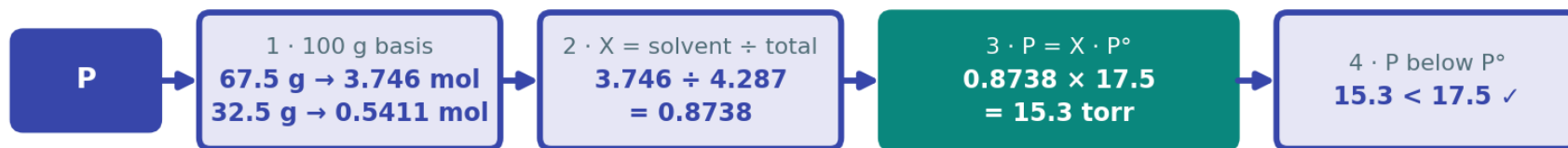
✓ Practice 5 answer: B

32.5% by mass: 32.5 g urea in every 100 g of solution

basis: take 100 g of solution · water = 100 – 32.5 = 67.5 g · 32.5 g urea → 0.5411 mol · 67.5 g water → 3.746 mol · $P^\circ = 17.5$ torr

$$X_{\text{water}} = \frac{3.746 \text{ mol water}}{(3.746 + 0.5411) \text{ mol}} = 0.8738 \rightarrow P = 0.8738 \times 17.5 = \mathbf{15.3 \text{ torr} \cdot \text{answer B}}$$

 **Sense check:** Urea is light, so 32.5 g is 0.541 mol, about an eighth of the moles. P falls by about an eighth: 17.5 – 15.3 = 2.2 torr. ✓





Check yourself

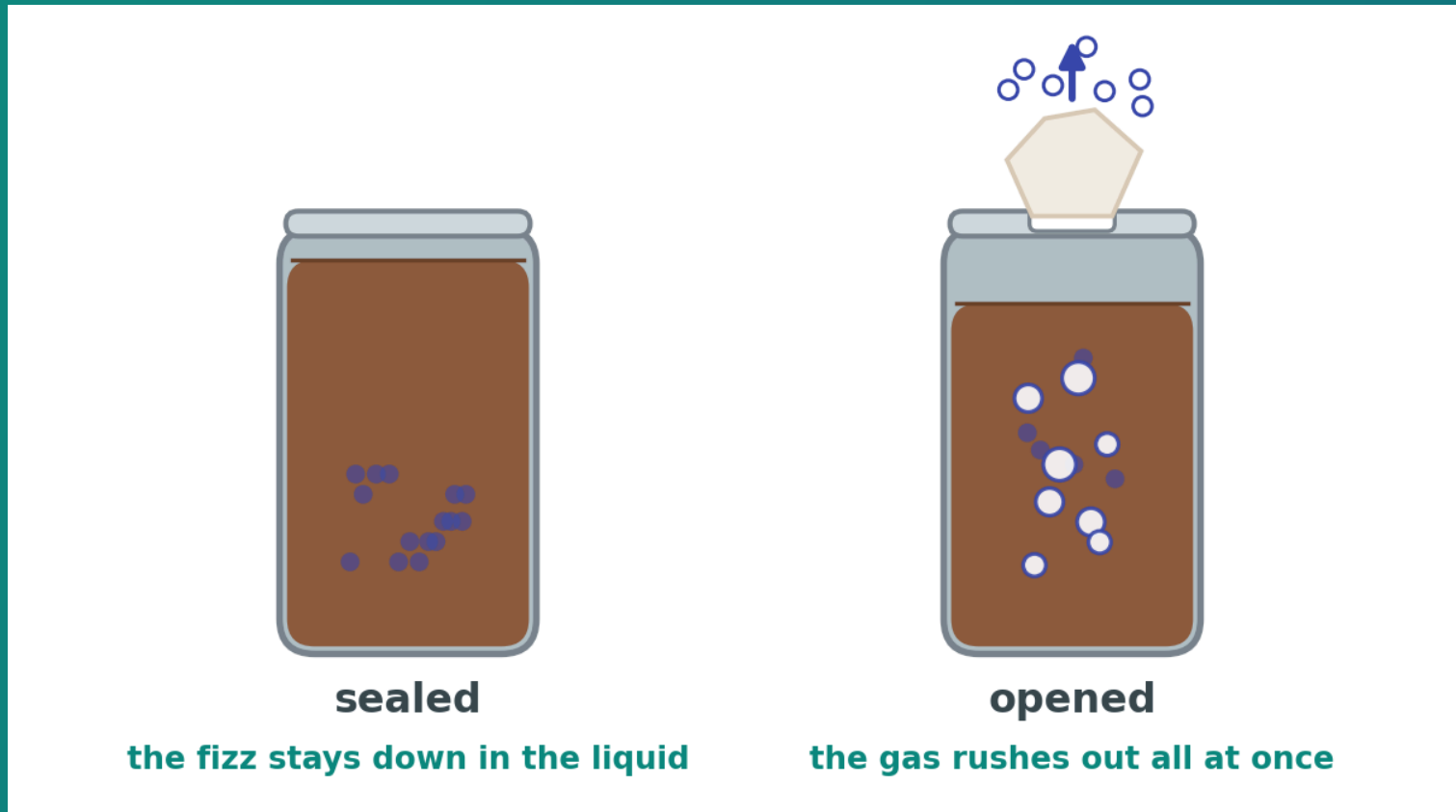
1. A solution holds 8.00 mol of water and 2.00 mol of a nonvolatile solute; pure water is 23.8 torr. Which mole fraction scales P° , and what is the solution's vapor pressure?
2. A nonvolatile solute drops a solvent's vapor pressure from 90.0 torr to 72.0 torr. Write the pressure lowering ΔP , and give the solvent's mole fraction.

A nonvolatile solute sets how much solvent escapes. A gas does the reverse from outside: Henry's law sets how much gas dissolves into a solvent, and that amount climbs with the gas's pressure above the liquid.

9 • 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



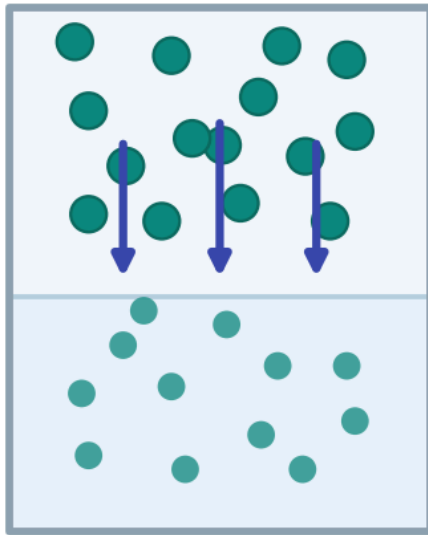
Crack the tab on a shaken can and it hisses, then foams over. Sealed, the fizz stayed down in the drink. Opened, the gas escapes all at once.

More gas overhead, more gas dissolved

more gas pressing on the surface → more gas dissolved

more gas overhead

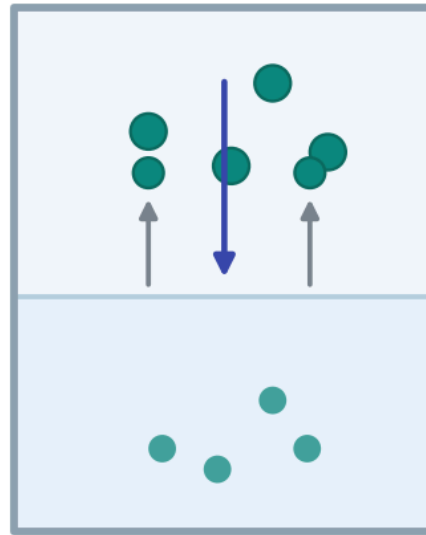
high pressure



much dissolves

less gas overhead

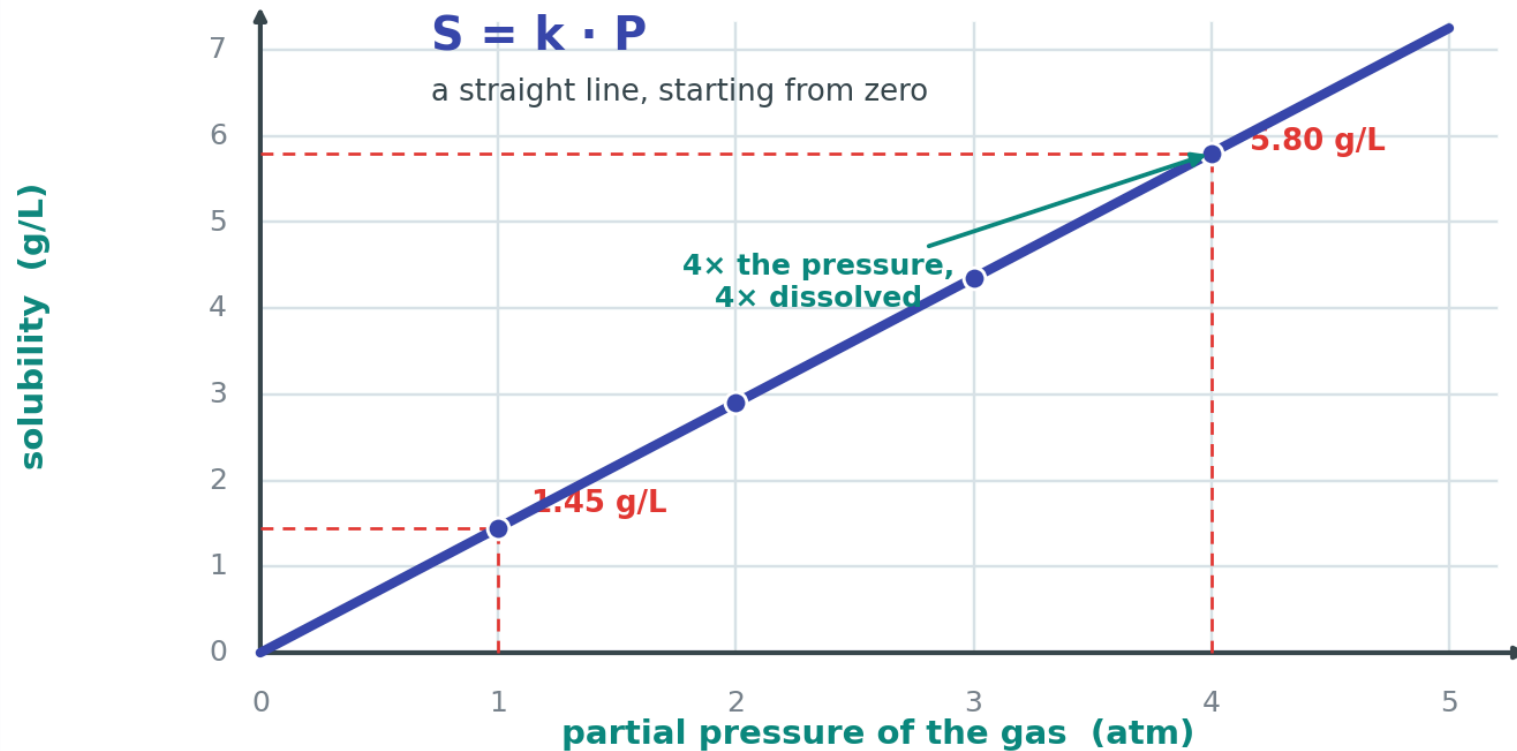
low pressure



little dissolves

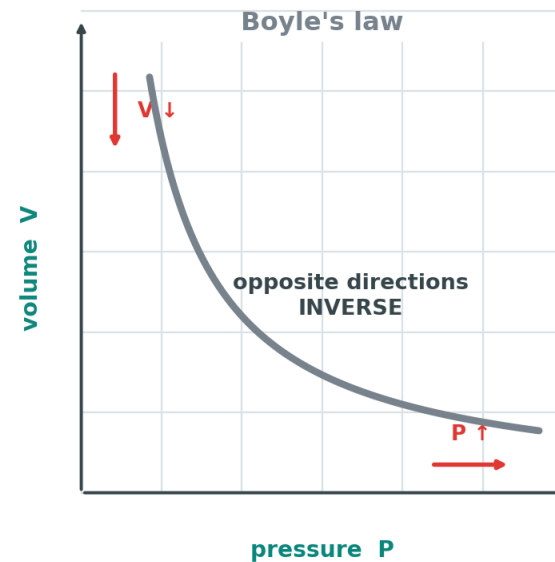
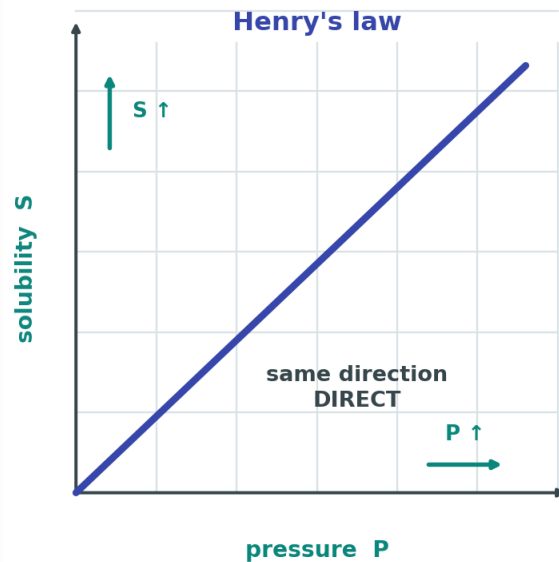
At constant temperature, a gas's solubility (how much dissolves) rises and falls in step with its partial pressure. More gas overhead drives more into solution.

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



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.



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

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

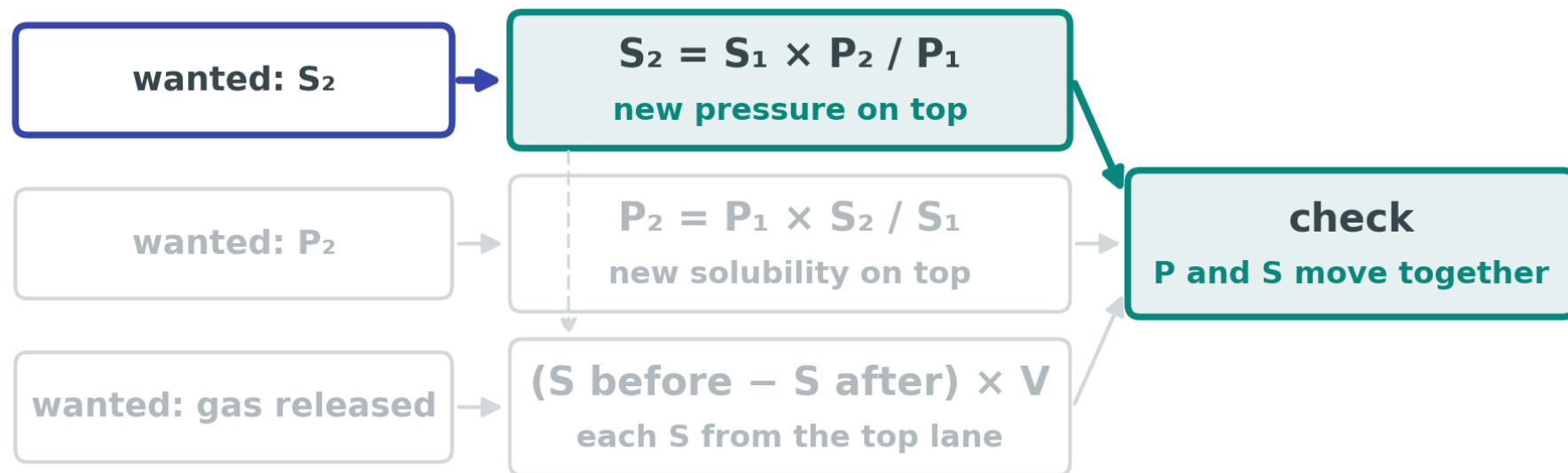
 **Sense check:** Pressure quadrupled at constant temperature, so four times as much CO_2 dissolves: $1.45 \rightarrow 5.80 \text{ g/L}$.

✓

✓ 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: the new pressure goes on top. 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

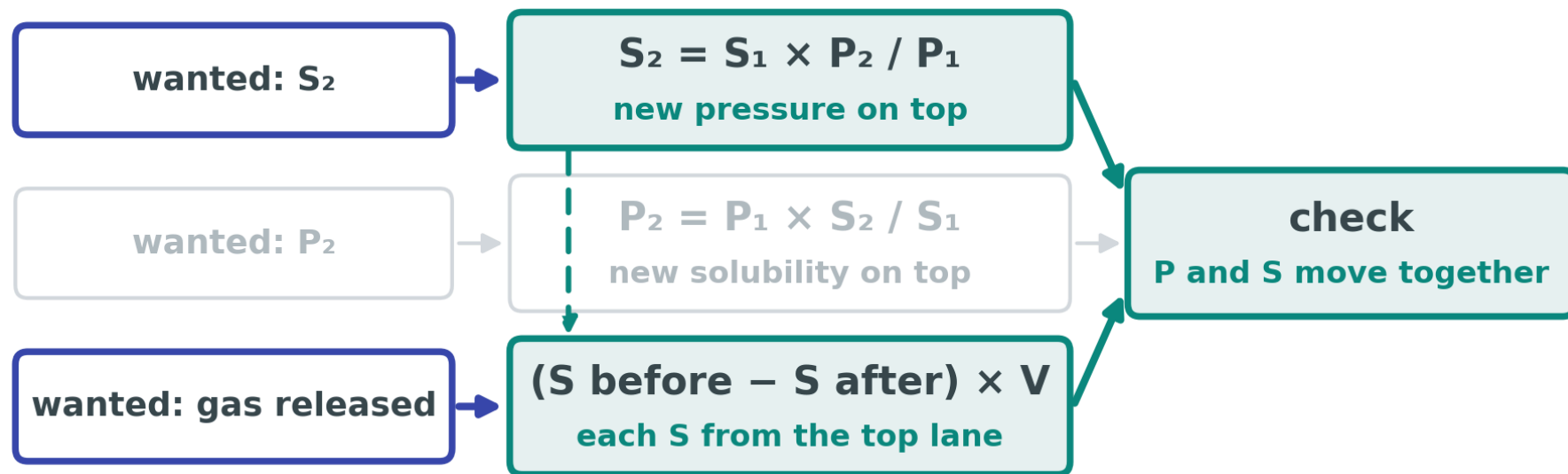
$$S_2 = 5.80 \text{ g/L} \times \frac{1.0 \text{ atm}}{4.0 \text{ atm}} = 1.45 \text{ g/L}$$

 **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 1

$$S_1/P_1 = S_2/P_2$$

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 1 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

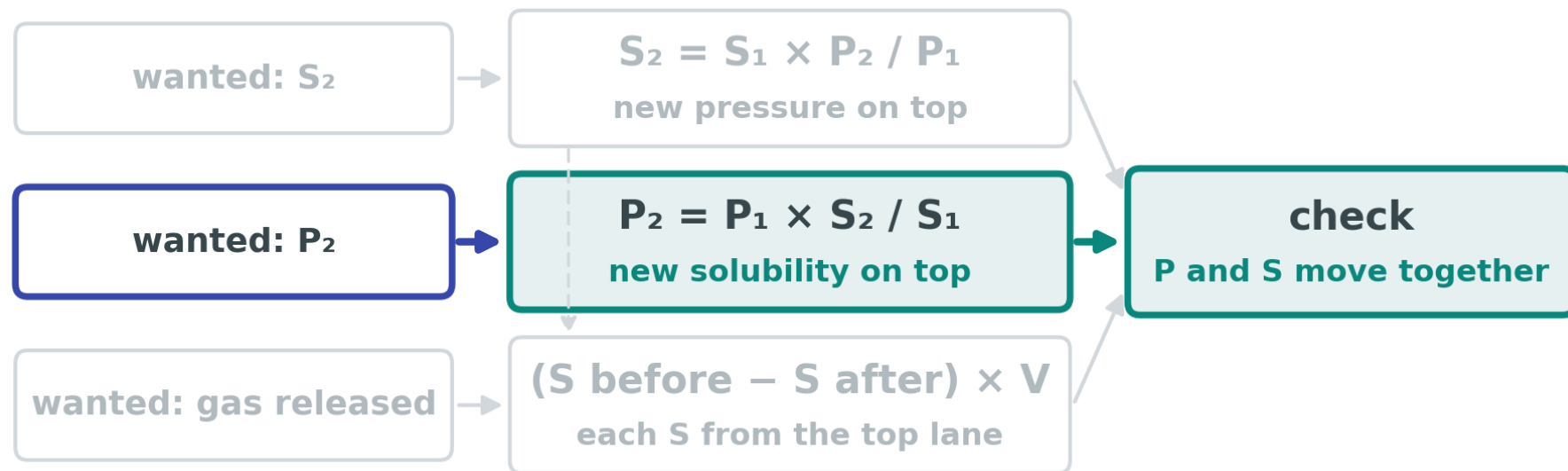
$$P_2 = 1.0 \text{ atm} \times \frac{8.70 \text{ g/L}}{1.45 \text{ g/L}} = \mathbf{6.0 \text{ atm}}$$

 **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 2

$$S_1/P_1 = S_2/P_2$$

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 2 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 3

$$S_1/P_1 = S_2/P_2$$

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 3 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 4

$$S_1/P_1 = S_2/P_2$$

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 4 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 5

$$S_1/P_1 = S_2/P_2 \cdot \text{CO}_2: 44.01 \text{ g/mol}$$

given: CO_2 0.0336 M at 1.00 atm · bottled under 3.50 atm · 591 mL · wanted: g CO_2 dissolved

At 25 °C, CO_2 dissolves to **0.0336 M** under 1.00 atm of CO_2 . A soda is bottled under **3.50 atm** of CO_2 at the same temperature. How many grams of CO_2 are dissolved in a **591 mL** bottle?

- A. 0.250
- B. 0.874
- C. 3.06
- D. 0.0695

✓ Practice 5 answer: C

$$S_1/P_1 = S_2/P_2 \cdot \text{CO}_2: 44.01 \text{ g/mol}$$

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

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

$$0.591 \text{ L} \times \frac{0.1176 \text{ mol}}{1 \text{ L}} \times \frac{44.01 \text{ g}}{1 \text{ mol}} = 3.06 \text{ g CO}_2 \cdot \text{answer C}$$

🧠 **Sense check:** 3.50 times the pressure dissolves 3.50 times the CO₂: 0.874 g × 3.50 = 3.06 g. ✓

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?

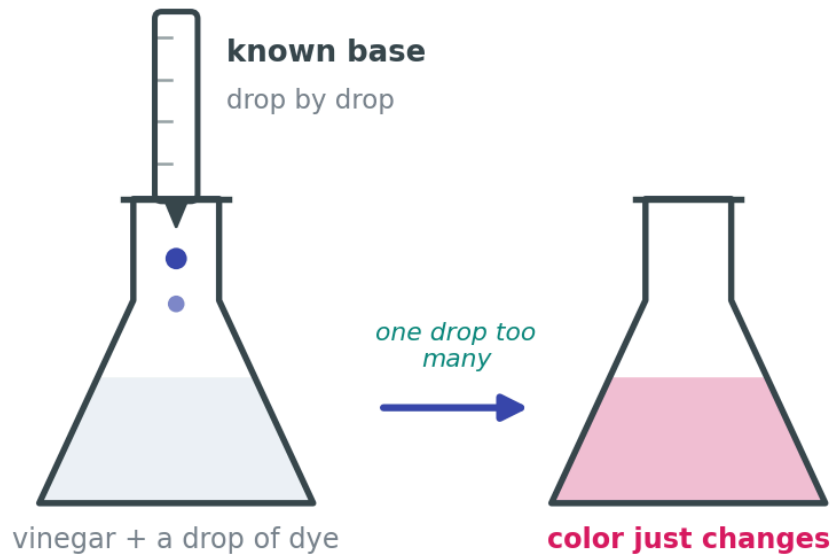
Henry's law reads a dissolved amount straight off a pressure. A titration reads an unknown concentration a different way: add a measured reactant until it exactly consumes the unknown, and the volume it took fixes the concentration.

10 • Titration Calculations

Find an unknown acid concentration from a titration by counting the titrant's millimoles, dividing out the balanced equation's base-to-acid ratio, and dividing by the acid's volume; the milliliters cancel.

Reading the acid in vinegar

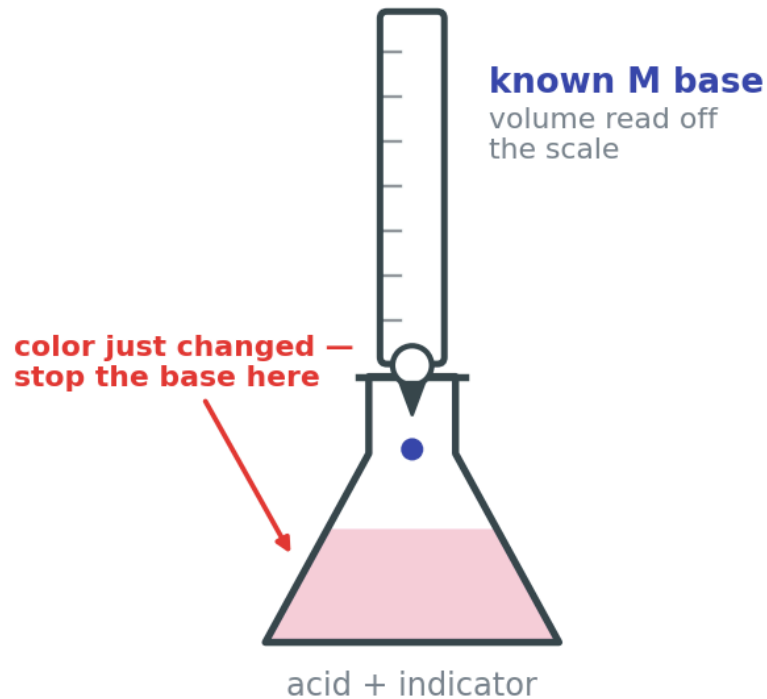
Count the base you added, and you know the acid.



Add a drop of dye to vinegar, then drip in a known base until the color just changes. The base you used measures the acid.

The equivalence point

The base exactly neutralizes the acid.



at the color change
moles OH^- added
=
moles H^+ available

Add a known base to an acid until an indicator flips; this addition is a titration. At the flip, the equivalence point, moles of OH^- added equal moles of H^+ available.

Equal millimoles in a 1:1 reaction

$$0.200 \text{ M} \times 25.00 \text{ mL} = 5.00 \text{ mmol}$$

molarity $\times$ volume, with volume in mL, counts millimoles

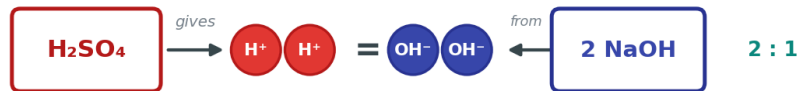
$$M_{\text{acid}} \cdot V_{\text{acid}} = M_{\text{base}} \cdot V_{\text{base}}$$

1 : 1 acid to base $\cdot$ both volumes in mL, so the mL cancel

Moles equal molarity times volume. With volume in milliliters, that product counts millimoles. When an acid and base react one-to-one, their millimoles are equal.

Polyprotic acids need the mole ratio

Each acidic H needs one OH⁻.



The balanced equation sets the base-to-acid ratio.

$$M_{\text{acid}} = M_{\text{base}} \cdot V_{\text{base}} \div (\text{ratio} \cdot V_{\text{acid}})$$

ratio = base units per acid unit · the mL cancel, so no liter step

A diprotic acid gives two H⁺ per unit; a triprotic gives three. Each H⁺ takes one OH⁻, so the balanced equation's coefficients set the base-to-acid ratio.

memory hook: count the H's written first

HCl 1 · H₂SO₄ 2 · H₃PO₄ 3 · that many NaOH per acid



The method

1. **Moles of the known solution:** $\text{molarity} \times \text{mL}$ gives millimoles.
2. **Apply the mole ratio:** cross to the other substance with the balanced equation.
3. **Divide to finish:** by its volume for a molarity, or by its molarity for a volume.



Worked example 1: a one-to-one titration



given: 20.00 mL HCl titrated by 25.00 mL of 0.200 M NaOH · wanted: M of the HCl

A **20.00 mL** sample of hydrochloric acid of unknown concentration is titrated with **0.200 M NaOH**. The indicator changes color after **25.00 mL** of base. Find the concentration of the acid.

Set it up: the base's millimoles, the 1:1 ratio, then divide by the acid's volume.

✓ Worked example 1: solution



given: 20.00 mL HCl · 25.00 mL of 0.200 M NaOH · wanted: M of the HCl

Step 1 · Moles of the known solution

$$25.00 \text{ mL base} \times \frac{0.200 \text{ mmol NaOH}}{1 \text{ mL base}} = 5.00 \text{ mmol NaOH}$$

Step 2 · Apply the mole ratio

Step 3 · Divide to finish

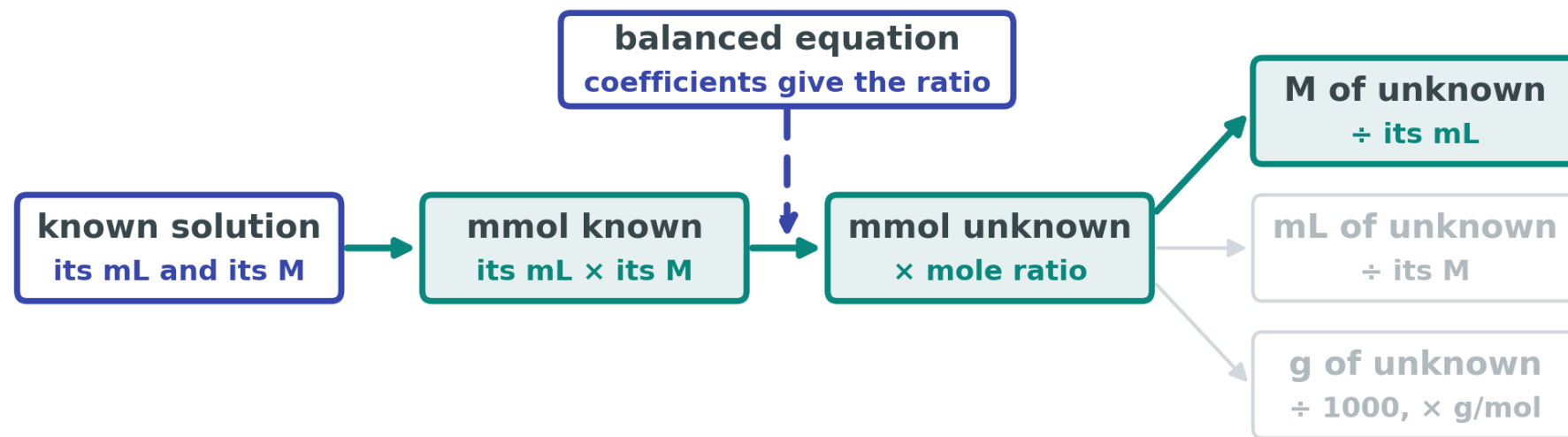
$$M = \frac{5.00 \text{ mmol HCl}}{20.00 \text{ mL acid}} = 0.250 \text{ M HCl}$$

 **Sense check:** The base's 25.00 mL slightly topped the acid's 20.00 mL at 0.200 M, so the acid runs a little higher: 0.250 M ✓

✓ Worked example 1: the route on the map



given: 25.00 mL of 0.200 M NaOH · 20.00 mL HCl · found: 0.250 M HCl



 **Sense check:** The base in the buret is the known solution. Its 5.00 mmol cross 1 : 1 to the acid, then the exit divides by the acid's own 20.00 mL. ✓



Worked example 2: a diprotic acid



given: 20.00 mL H_2SO_4 titrated by 32.00 mL of 0.250 M NaOH · wanted: M of the H_2SO_4

A **20.00 mL** sample of sulfuric acid is titrated with **0.250 M NaOH**, and the indicator changes color after **32.00 mL** of base. Sulfuric acid is diprotic: each unit gives two H^+ . Find its concentration.

A tempting shortcut: carry the base's millimoles straight to the acid's volume. Test it against the balanced equation.

✓ Worked example 2: solution



given: 20.00 mL H_2SO_4 · 32.00 mL of 0.250 M NaOH · wanted: M of the H_2SO_4

Step 1 · Moles of the known solution

The tempting shortcut

Carry the base's millimoles straight to the acid's volume, as if the ratio were one:

$$M = \frac{8.00 \text{ mmol NaOH}}{20.00 \text{ mL acid}} = 0.400 \text{ M} \times$$

 **Sense check:** Each H_2SO_4 gives two H^+ , so those 8.00 mmol OH^- neutralized only half as many acid units. ✗

✓ Worked example 2: the mole ratio



given: 20.00 mL H_2SO_4 · 32.00 mL of 0.250 M NaOH · wanted: M of the H_2SO_4

Step 2 · Apply the mole ratio

Step 3 · Divide to finish

$$32.00 \text{ mL base} \times \frac{0.250 \text{ mmol NaOH}}{1 \text{ mL base}} \times \frac{1 \text{ mmol H}_2\text{SO}_4}{2 \text{ mmol NaOH}}$$

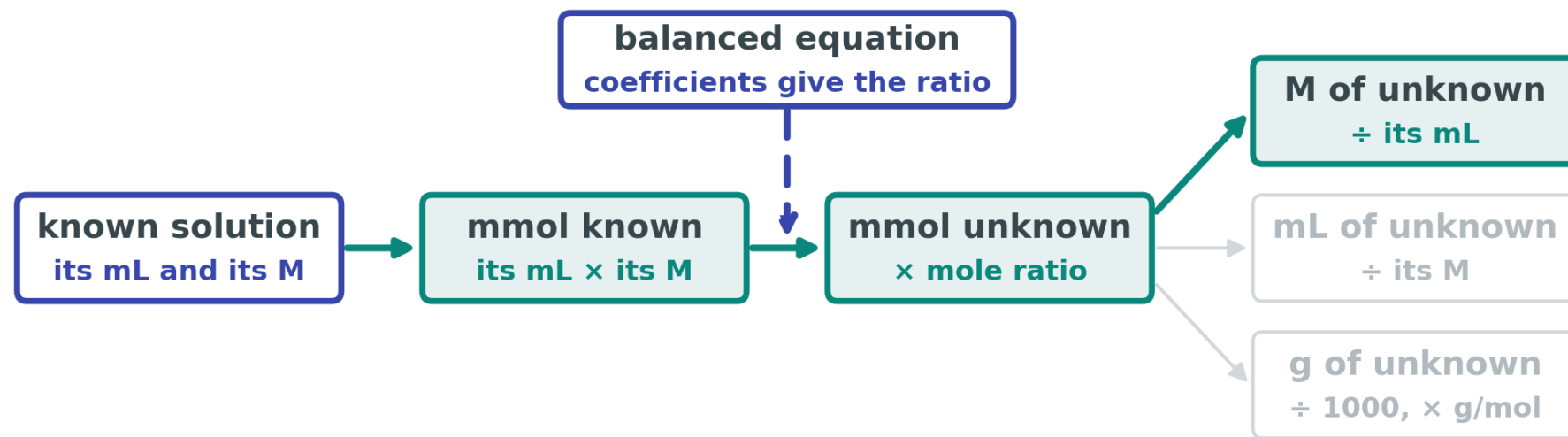
= 4.00 mmol H_2SO_4 , then $\div$ 20.00 mL acid = **0.200 M H_2SO_4**

 **Sense check:** The 2:1 ratio halves the acid's millimoles, so its concentration is half the shortcut's guess: 0.200 M, not 0.400 M ✓

✓ Worked example 2: the route on the map



given: 32.00 mL of 0.250 M NaOH · 20.00 mL H₂SO₄ · found: 0.200 M H₂SO₄



 **Sense check:** The same route as a 1 : 1 titration. Only the ratio step changes: 1 H₂SO₄ per 2 NaOH halves 8.00 mmol to 4.00 mmol. ✓

Take-home: a polyprotic acid is not 1:1



given: 32.00 mL of 0.250 M NaOH · 20.00 mL H₂SO₄

$$M = \frac{8.00 \text{ mmol NaOH}}{20.00 \text{ mL acid}} = 0.400 \text{ M} \times \cdot \text{counts each H}_2\text{SO}_4 \text{ as one H}^+$$

$$32.00 \text{ mL base} \times \frac{0.250 \text{ mmol NaOH}}{1 \text{ mL base}} \times \frac{1 \text{ mmol H}_2\text{SO}_4}{2 \text{ mmol NaOH}} \div 20.00 \text{ mL acid} = \mathbf{0.200 \text{ M} \checkmark}$$

A diprotic acid feeds two H⁺, a triprotic three. The balanced equation's ratio divides the base's millimoles before the volume does. Skip it and every diprotic answer comes out doubled.

Your turn: sulfuric acid and KOH



given: 40.00 mL of 0.150 M KOH · 25.00 mL H_2SO_4 · wanted: M of the H_2SO_4

It takes **40.00 mL of 0.150 M KOH** to titrate **25.00 mL of H_2SO_4** . Fill the molarity, the 2:1 ratio, and the acid's volume, then compute:

$$40.00 \text{ mL base} \times \frac{\text{_____ mmol KOH}}{1 \text{ mL base}} \times \frac{1 \text{ mmol H}_2\text{SO}_4}{\text{_____ mmol KOH}}$$

$$= \text{_____ mmol H}_2\text{SO}_4, \text{ then } \div \text{_____ mL acid} = \text{_____ M}$$

$$40.00 \text{ mL base} \times \frac{0.150 \text{ mmol KOH}}{1 \text{ mL base}} \times \frac{1 \text{ mmol H}_2\text{SO}_4}{2 \text{ mmol KOH}}$$

$$= 3.00 \text{ mmol H}_2\text{SO}_4, \text{ then } \div 25.00 \text{ mL acid} = \mathbf{0.120 \text{ M H}_2\text{SO}_4}$$

⚠ Where this goes wrong



given: 32.00 mL of 0.250 M NaOH · 20.00 mL H₂SO₄ · correct: 8.00 mmol NaOH → 4.00 mmol H₂SO₄ → 0.200 M

Treating the diprotic acid as 1:1. 8.00 mmol ÷ 20.00 mL = 0.400 M, double the truth. Each H₂SO₄ gives two H⁺, so divide the base's millimoles by 2 first.

Multiplying by the ratio instead of dividing. 2 × 8.00 ÷ 20.00 = 0.800 M. Each acid unit uses up two OH⁻, so there are fewer acid units than base units: the 2 divides.

Swapping the two volumes. Pairing 0.250 M with 20.00 mL gives 0.250 × 20.00 ÷ 2 ÷ 32.00 = 0.0781 M. Each molarity multiplies its own solution's volume.

 Practice 1

titrant 0.125 M NaOH · sample 18.00 mL oxalic acid

It takes **28.80 mL of 0.125 M NaOH** to reach the color change while titrating **18.00 mL of oxalic acid**. What is the molarity of the oxalic acid?

- A. 1.80
- B. 0.100
- C. 0.200
- D. 0.400

✓ Practice 1 answer: B



given: 18.00 mL $\text{H}_2\text{C}_2\text{O}_4$ · 28.80 mL of 0.125 M NaOH · wanted: M of the $\text{H}_2\text{C}_2\text{O}_4$

$$28.80 \text{ mL base} \times \frac{0.125 \text{ mmol NaOH}}{1 \text{ mL base}} \times \frac{1 \text{ mmol H}_2\text{C}_2\text{O}_4}{2 \text{ mmol NaOH}}$$

= 1.80 mmol $\text{H}_2\text{C}_2\text{O}_4$, then $\div$ 18.00 mL acid = **0.100 M · answer B**

A stopped at 1.80 mmol of oxalic acid, before dividing by its 18.00 mL. C ignored the 2:1 ratio: $3.60 \div 18.00 = 0.200 \text{ M}$. D multiplied by the ratio: $2 \times 3.60 \div 18.00 = 0.400 \text{ M}$.

 **Sense check:** Two NaOH per oxalic acid, so the acid holds half the base's 3.60 mmol: 1.80 mmol in 18.00 mL is 0.100 M ✓



Worked example 3: volume of base for a triprotic acid



given: 25.00 mL of 0.0600 M H_3PO_4 · titrant 0.150 M NaOH · wanted: mL of NaOH

A **25.00 mL** sample of **0.0600 M phosphoric acid** is titrated with **0.150 M NaOH**.

Phosphoric acid is triprotic: each unit gives three H^+ . What volume of base reaches the color change?

A common first attempt divides by 3 out of habit. Test it against the protons.

✓ Worked example 3: solution



given: 25.00 mL of 0.0600 M H_3PO_4 · 0.150 M NaOH · wanted: mL of NaOH

Step 1 · Moles of the known solution

A common first attempt

$$1.50 \text{ mmol} \div 3 = 0.500 \text{ mmol NaOH} \rightarrow 0.500 \div 0.150 = 3.33 \text{ mL} \times$$

 **Sense check:** Each H_3PO_4 carries three H^+ , and each H^+ takes one OH^- . The base's millimoles are three times the acid's, not a third. $\times$

✓ Worked example 3: the mole ratio



given: 25.00 mL of 0.0600 M H_3PO_4 · 0.150 M NaOH · wanted: mL of NaOH

Step 2 · Apply the mole ratio

Step 3 · Divide to finish

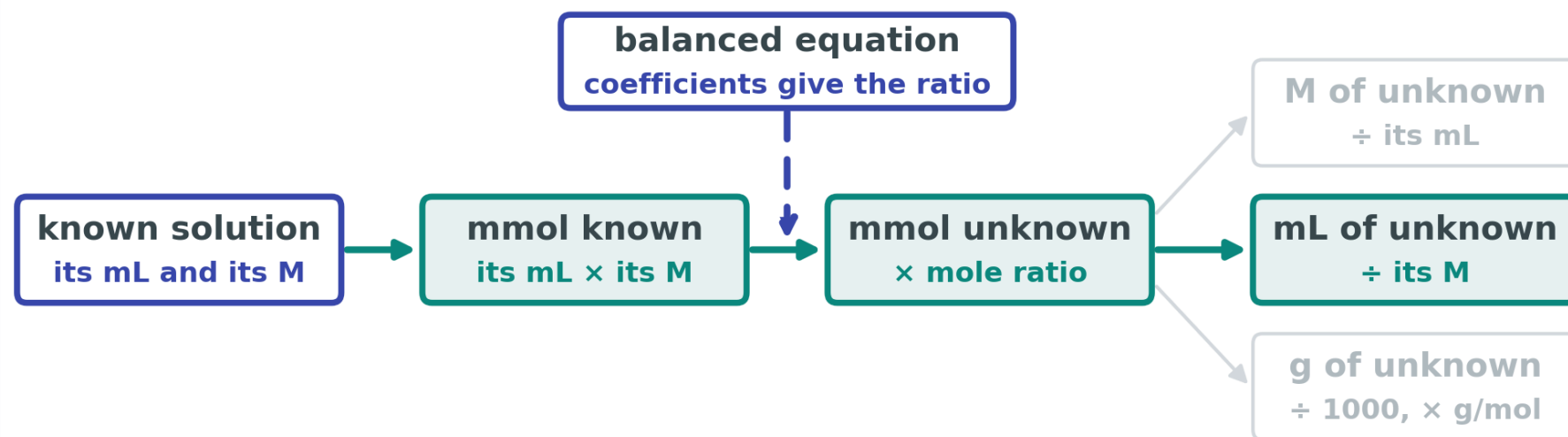
$$1.50 \text{ mmol } \text{H}_3\text{PO}_4 \times \frac{3 \text{ mmol NaOH}}{1 \text{ mmol } \text{H}_3\text{PO}_4} \times \frac{1 \text{ mL base}}{0.150 \text{ mmol NaOH}} = 30.0 \text{ mL NaOH}$$

🧠 **Sense check:** 4.50 mmol of NaOH at 0.150 mmol per mL takes 30.0 mL, nine times the 3.33 mL guess ✓

✓ Worked example 3: the route on the map



given: 25.00 mL of 0.0600 M H_3PO_4 · 0.150 M NaOH · found: 30.0 mL NaOH



 **Sense check:** The acid is the known solution. The ratio triples 1.50 mmol to 4.50 mmol NaOH, and the volume exit divides by the base's 0.150 M. ✓



Practice 2



given: 26.40 mL of 0.120 M NaOH · citric acid 192.12 g/mol · wanted: g of citric acid

A powdered drink mix is dissolved in water, and its citric acid takes **26.40 mL of 0.120 M NaOH** to reach the color change. How many grams of citric acid did the sample hold?

- A. 0.203
- B. 0.609
- C. 1.06
- D. 203
- E. 1.83

✓ Practice 2 answer: A



26.40 mL of 0.120 M NaOH · citric acid 192.12 g/mol

$$26.40 \text{ mL base} \times \frac{0.120 \text{ mmol NaOH}}{1 \text{ mL base}} \times \frac{1 \text{ mmol acid}}{3 \text{ mmol NaOH}} = 1.056 \text{ mmol acid}$$

$$1.056 \text{ mmol} \times \frac{1 \text{ mol}}{1000 \text{ mmol}} \times \frac{192.12 \text{ g}}{1 \text{ mol}} = \mathbf{0.203 \text{ g} \cdot \text{answer A}}$$

B ignored the 3:1 ratio: 3.168 mmol → 0.609 g. C stopped at 1.06 mmol, before the molar mass. D skipped mmol → mol: 1.056 × 192.12 = 203 g. E multiplied by the ratio: 9.504 mmol → 1.83 g.

 **Sense check:** 1.056 mmol is a third of 3.168: three NaOH per citric acid. 203 g would outweigh the sample. ✓



Extra practice 1



titrant KOH, molarity unknown · sample 10.00 mL of 0.0450 M H_3PO_4

A new bottle of KOH solution needs **12.50 mL** to neutralize **10.00 mL of 0.0450 M H_3PO_4** .
What is the molarity of the KOH?

- A. 0.0360
- B. 0.0120
- C. 0.108
- D. 0.0600

✓ Extra practice 1 answer: C



given: 10.00 mL of 0.0450 M H_3PO_4 · 12.50 mL KOH · wanted: M of the KOH

$$10.00 \text{ mL acid} \times \frac{0.0450 \text{ mmol H}_3\text{PO}_4}{1 \text{ mL acid}} \times \frac{3 \text{ mmol KOH}}{1 \text{ mmol H}_3\text{PO}_4}$$

= 1.35 mmol KOH, then $\div$ 12.50 mL base = **0.108 M · answer C**

A ignored the 3 : 1 ratio: $0.450 \div 12.50 = 0.0360 \text{ M}$. B divided by the ratio: $0.450 \div 3 \div 12.50 = 0.0120 \text{ M}$. D divided by the total volume: $1.35 \div 22.50 = 0.0600 \text{ M}$.

 **Sense check:** The acid is the known solution. Each H_3PO_4 takes three KOH, so the base's millimoles are triple the acid's 0.450 mmol: 1.35 mmol in 12.50 mL is 0.108 M ✓



Extra practice 2



sample 0.2933 g tartaric acid, 150.09 g/mol · titrant NaOH, molarity unknown

A **0.2933 g** sample of solid tartaric acid is dissolved in water and titrated with NaOH. The color change comes after **23.90 mL** of base. What is the molarity of the NaOH?

- A. 0.6136
- B. 0.08176
- C. 3.908
- D. 0.04088
- E. 0.1635

✓ Extra practice 2 answer: E



given: 0.2933 g tartaric acid, 150.09 g/mol · 23.90 mL NaOH · wanted: M of the NaOH

$$0.2933 \text{ g} \times \frac{1 \text{ mol}}{150.09 \text{ g}} \times \frac{1000 \text{ mmol acid}}{1 \text{ mol}} \times \frac{2 \text{ mmol NaOH}}{1 \text{ mmol acid}}$$

= 3.908 mmol NaOH, then ÷ 23.90 mL base = **0.1635 M · answer E**

A used NaOH's molar mass: $0.2933 \div 40.00 \times 1000 \times 2 \div 23.90 = 0.6136 \text{ M}$. B ignored the 2 : 1 ratio: $1.954 \div 23.90 = 0.08176 \text{ M}$. C stopped at 3.908 mmol NaOH, before dividing by its volume. D divided by the ratio: $1.954 \div 2 \div 23.90 = 0.04088 \text{ M}$.

 **Sense check:** Two NaOH per tartaric acid: 3.908 mmol of base, double the acid's 1.954 mmol ✓



Extra practice 3



stock 5.00 mL diluted to 100.0 mL · titrated 10.00 mL portion · titrant 0.1000 M NaOH

A **5.00 mL** sample of a nitric acid stock is diluted with water to **100.0 mL**. A **10.00 mL** portion of the diluted acid takes **17.40 mL of 0.1000 M NaOH** to reach the end point. What is the molarity of the stock?

- A. 0.174
- B. 0.00870
- C. 1.15
- D. 3.48
- E. 1.74

✓ Extra practice 3 answer: D



given: 5.00 mL stock → 100.0 mL · 10.00 mL portion · 17.40 mL of 0.1000 M NaOH · wanted: M_1

$$17.40 \text{ mL base} \times \frac{0.1000 \text{ mmol NaOH}}{1 \text{ mL base}} \times \frac{1 \text{ mmol HNO}_3}{1 \text{ mmol NaOH}}$$

$$= 1.740 \text{ mmol HNO}_3 \div 10.00 \text{ mL} = 0.1740 \text{ M diluted} \cdot M_1 = 0.1740 \text{ M} \times 100.0 \text{ mL} \div 5.00 \text{ mL}$$

$$= \mathbf{3.48 \text{ M} \cdot \text{answer D}}$$

A stopped at the diluted acid: 0.174 M. B ran the dilution backward: $0.1740 \times 5.00 \div 100.0 = 0.00870 \text{ M}$. C swapped the titration volumes: $0.05747 \text{ M} \times 20 = 1.15 \text{ M}$. E used the 10.00 mL portion as V_1 : $0.1740 \times 100.0 \div 10.00 = 1.74 \text{ M}$.

 **Sense check:** The stock was diluted 20-fold, so it is 20 times the titrated acid: 3.48 M ✓

Check yourself

1. A 1:1 titration takes **24.0 mL of 0.150 M NaOH** to neutralize **20.0 mL of HCl**. Set up $M_{\text{acid}} = M_{\text{base}} \cdot V_{\text{base}} \div V_{\text{acid}}$ and find the HCl concentration.
2. Phosphoric acid reacts as $\text{H}_3\text{PO}_4 + 3 \text{NaOH} \rightarrow \text{Na}_3\text{PO}_4 + 3 \text{H}_2\text{O}$. Which number becomes the base-to-acid ratio in the denominator, and where does it come from?

Every calculation here counted moles with a concentration, then crossed substances with the balanced equation's ratio: the same mole-ratio bridge that drives reaction stoichiometry. A concentration is only another way to count the moles a reaction runs on.

Can you...?

- ☐ explain "like dissolves like" and predict whether a solute dissolves in a given solvent?
- ☐ describe unsaturated, saturated, and supersaturated solutions; read a solubility curve?
- ☐ use the dilution relation $M_1V_1 = M_2V_2$ to find a concentration or volume?
- ☐ calculate molality and mole fraction, and convert among molarity, molality, and mass percent?
- ☐ classify a mixture as a solution, a colloid, or a suspension from particle size, light scattering, and settling, and name a colloid's type from its dispersed phase and medium?
- ☐ apply colligative properties (freezing-point depression, boiling-point elevation, and Raoult's law) using the van't Hoff factor?
- ☐ apply Henry's law to gas solubility and titration stoichiometry to find an unknown concentration?

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