

Foundations & Measurement

CHEM 50 · Dr. Karmach

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

- Define matter, place a description in the macroscopic, microscopic, or symbolic domain, and classify a statement as hypothesis, law, or theory
- Name the SI base units, tell a base unit from a derived unit, and pick a sensible metric unit for a measurement
- Classify matter as an element, a compound, or a mixture
- Distinguish precision from accuracy in measured data
- Report measurements and results with the correct significant figures
- Convert units with conversion factors, canceling units at each step
- Use density as a conversion factor between mass and volume

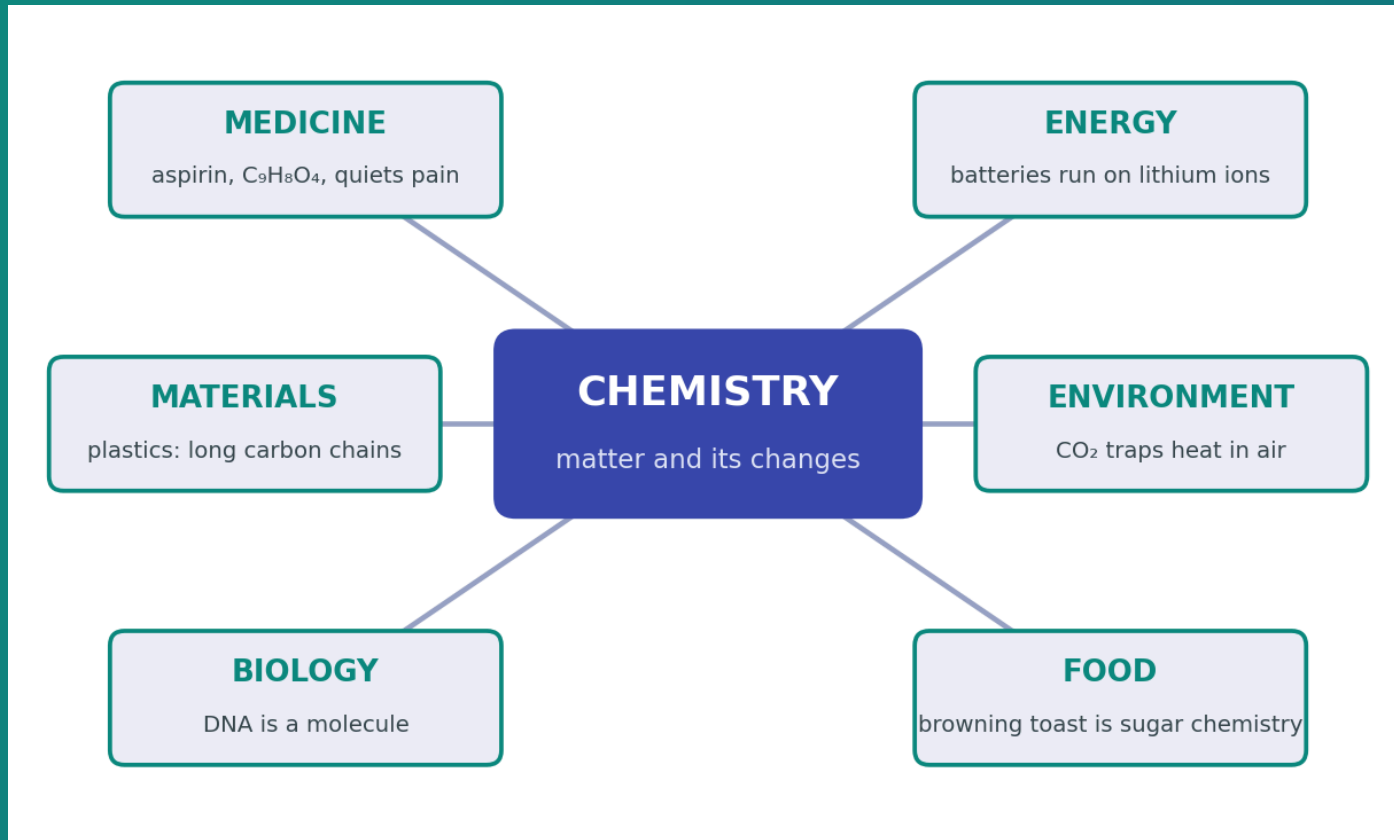
Today's route

1. **Chemistry in Context**
2. **Classifying Matter**
3. **Physical & Chemical Properties**
4. **States of Matter**
5. **Precision & Accuracy**
6. **Scientific Notation**
7. **Significant Figures**
8. **The SI Unit System**
9. **Dimensional Analysis**
10. **Temperature Scales & Conversions**
11. **Density as a Conversion Factor**

1 • Chemistry in Context

Define matter, place any description in the macroscopic, microscopic, or symbolic domain, and classify a scientific statement as hypothesis, law, or theory.

Matter changes all day long



Striking a match, charging a phone, digesting lunch: each is matter changing. One science studies matter and its changes, so medicine, farming, and engineering all lean on it.

The science of matter

Matter is anything that has mass and occupies space. Chemistry studies matter and its changes, and every science that handles material draws on it: the central science.

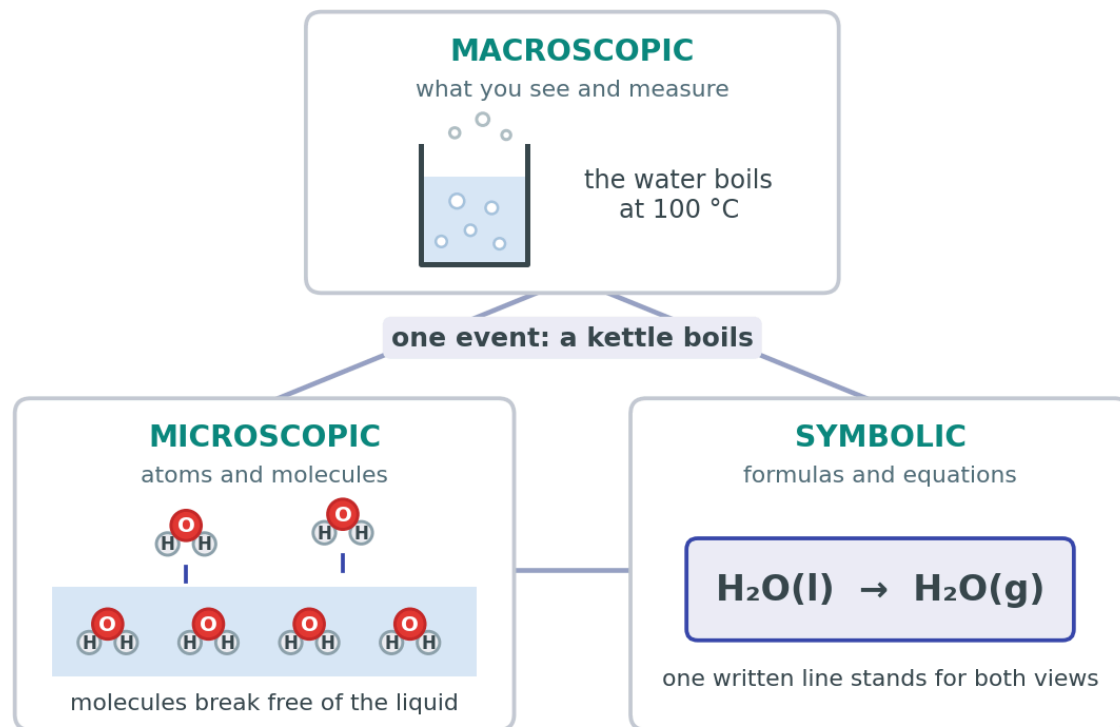
matter: has mass, occupies space

air, rust, salt, seawater: all matter

not matter: light, heat, an idea

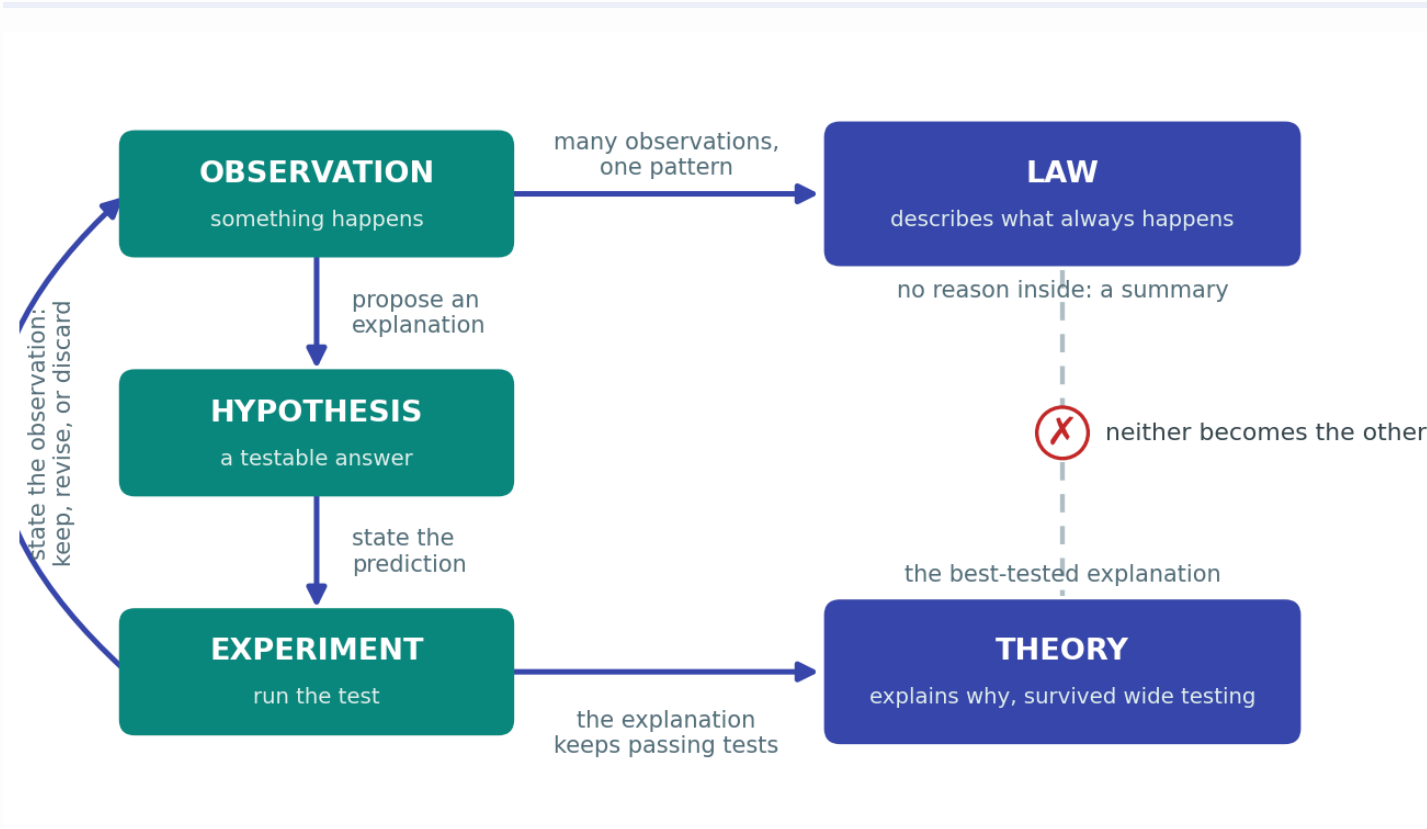
no mass, no space filled: energy and information

The three domains



Chemists work one event in three domains. The macroscopic domain holds what you see and measure. The microscopic domain holds atoms and molecules. The symbolic domain holds the formulas and equations that stand for both.

The scientific method



An observation raises a question. A hypothesis proposes a testable answer. An experiment tests it: state the prediction, state the observation, keep or revise. Every answer invites a sharper test.

Hypothesis, law, theory

A hypothesis proposes a testable answer, still on trial. A law summarizes what always happens. A theory explains why, and has survived wide testing. Laws describe; theories explain.

hypothesis: the pond turned green because fertilizer ran in
proposed and testable: sample the water upstream and down

law: every gas held at steady pressure expands when heated
the summary: centuries of thermometers and pistons agree

theory: hot particles move faster and hit harder
the tested explanation standing behind the law

Check yourself

1. Three descriptions of one fizzing tablet: the tablet visibly shrinks; CO_2 molecules leave the liquid; $\text{CO}_2(\text{aq}) \rightarrow \text{CO}_2(\text{g})$. Place each in its domain.
2. State conservation of mass twice: once as the law, once as the theory that explains it.
Could more testing ever turn one into the other?

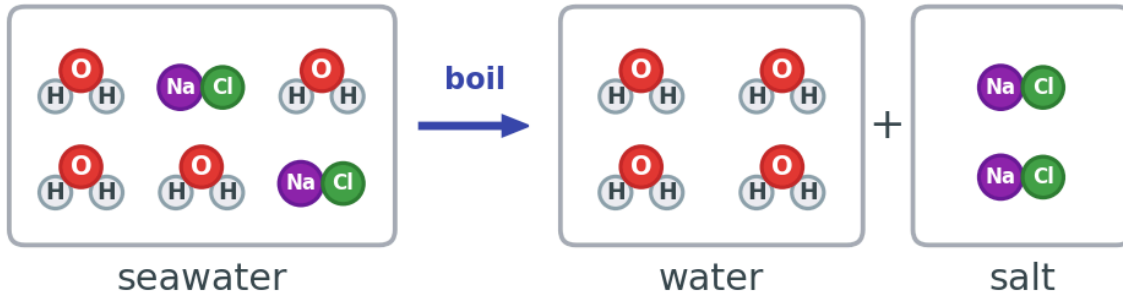
The scientific method built chemistry's biggest ideas: experiments forced the model of the atom itself to be redrawn, twice. And the first job the method hands a chemist is sorting matter, because every sample is an element, a compound, or a mixture, and two counts place it.

2 · Classifying Matter

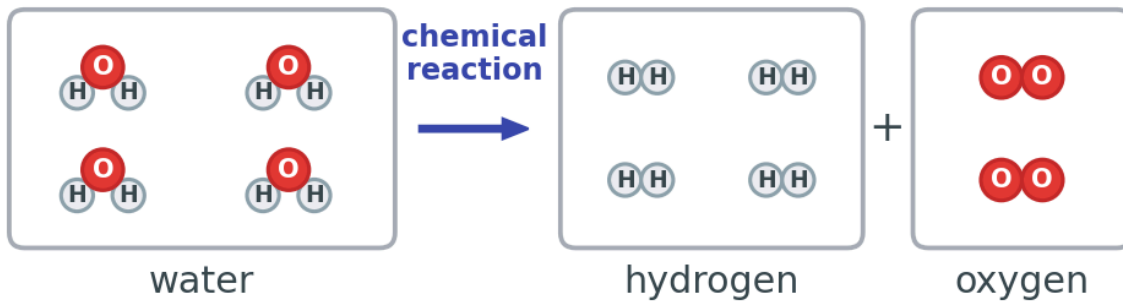
Classify any sample as an element, a compound, or a homogeneous or heterogeneous mixture by counting kinds of particles and phases.

Two ways to take matter apart

boiling separates seawater into water and salt



splitting water takes a chemical reaction

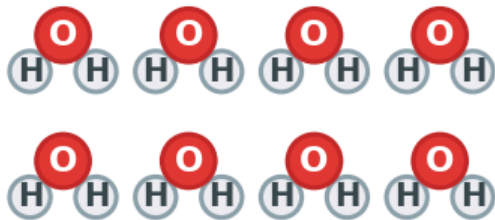


Boiling seawater leaves salt behind. Splitting water takes a chemical reaction. A sample's class tells what can pull it apart, and whether a fixed formula exists.

Pure substance or mixture

pure substance

one kind of particle

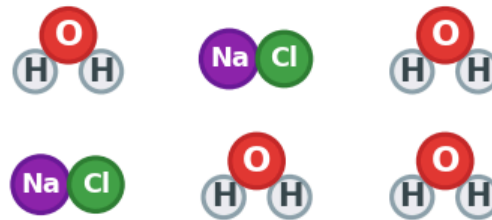


water

every particle identical: composition fixed

mixture

two or more kinds of particles



salt water

more salt or less: composition varies

A **pure substance** contains one kind of particle, so its composition is fixed. A **mixture** contains two or more kinds; its composition can vary, and each component keeps its identity.

Element or compound

Both are pure substances. An **element** contains one kind of atom; no chemical reaction breaks it down. A **compound** contains two or more elements bonded in a fixed ratio; a chemical reaction can take it apart.

elements: Cu · O₂ · S₈

one kind of atom each: 118 elements known

compounds: H₂O · NaCl · CO₂

two or more elements, in a ratio that never changes

Homogeneous or heterogeneous

A **phase** is a uniform region separated from its neighbors by a physical boundary. One phase throughout: homogeneous. Two or more phases: heterogeneous.

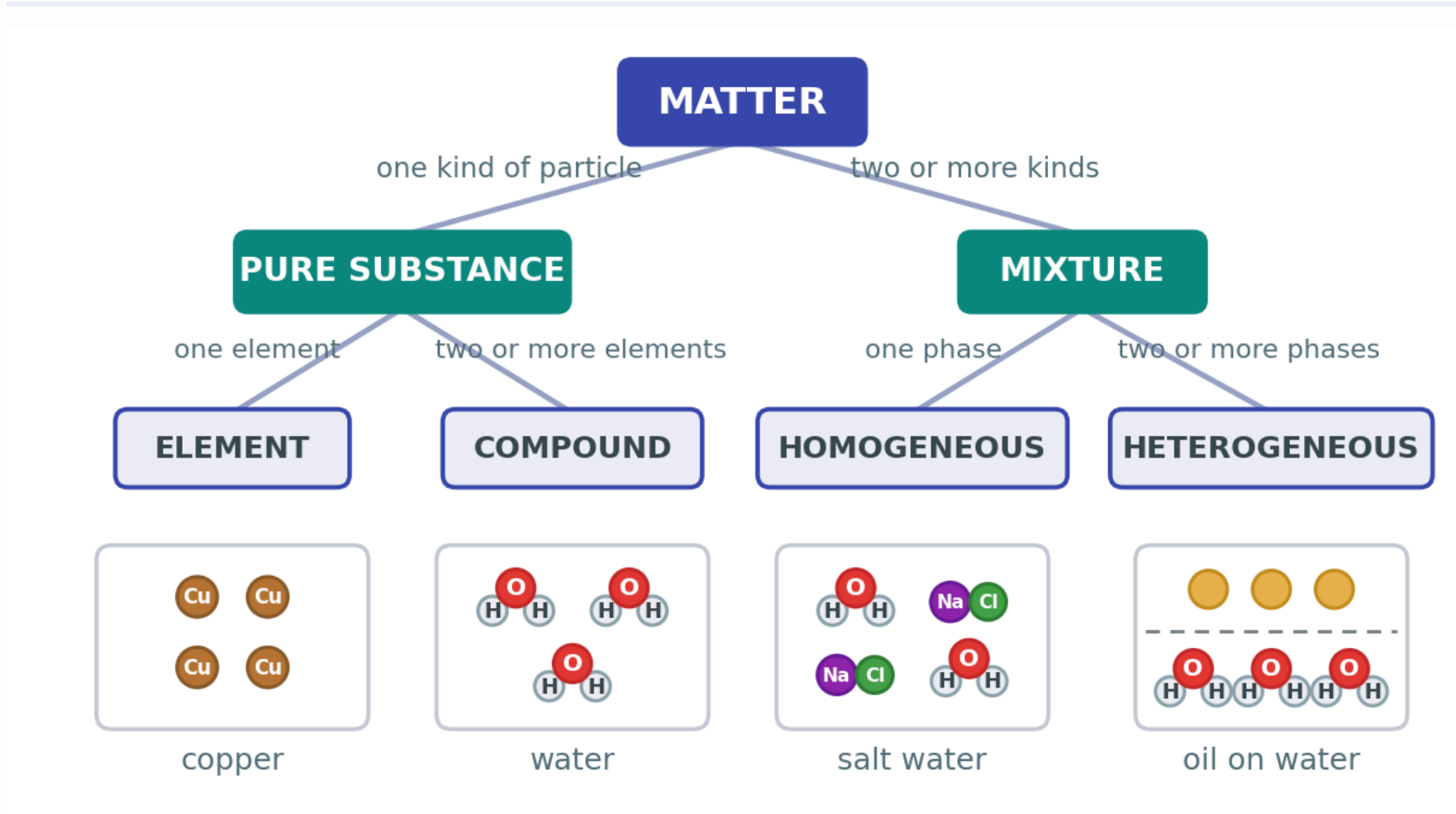
salt water: one phase

homogeneous: any drop matches any other drop

oil on water: two phases

heterogeneous: the top sample is oil, the bottom sample is water

The complete map



Two counts place any sample: the kinds of particles, then the elements or the phases.



The method

1. **Count the kinds of particles.** One → pure substance; more → mixture.
2. **Count the elements or the phases.** Pure: elements. Mixture: phases.
3. **Name the class.** One element → element. More → compound. One phase → homogeneous. More → heterogeneous.



Worked example 1: carbon dioxide

CO₂ from a fire extinguisher

every particle: 1 C + 2 O = 3 atoms, the same in every molecule

A CO₂ fire extinguisher discharges nothing but carbon dioxide. Classify the gas.

✓ Worked example 1: solution

CO₂ from a fire extinguisher

every particle: 1 C + 2 O = 3 atoms, the same in every molecule

Step 1 · Count the kinds of particles

Step 2 · Count the elements or the phases

Step 3 · Name the class

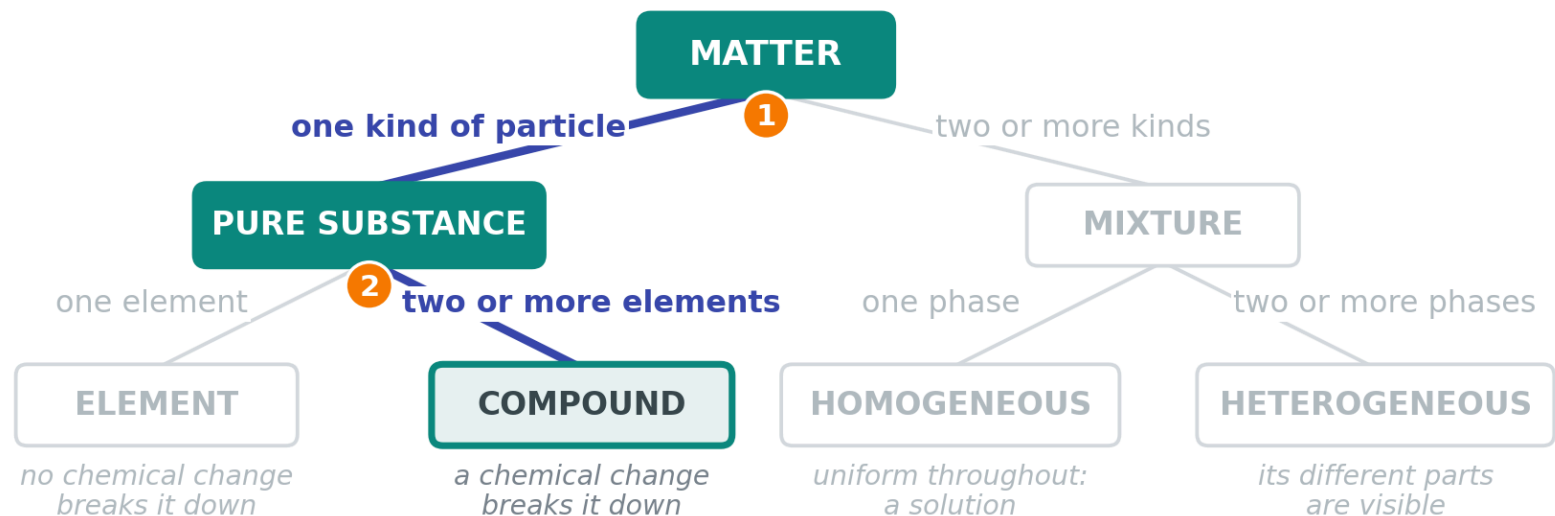
CO₂ → a compound

pure substance · two elements · fixed 1 C : 2 O ratio

 **Sense check:** Every CO₂ molecule carries the same 1 : 2 ratio, and only a chemical reaction separates the carbon from the oxygen. Fixed composition marks a compound.

✓ Worked example 1: the route on the map

CO₂ from a fire extinguisher
 given: every particle is CO₂ · found: a compound



① count the kinds of particles ② count the elements (pure) or the phases (mixture) ③ name the class

Sense check: Count 1: one kind of particle → pure substance. Count 2: two elements, C and O → compound. ✓



Worked example 2: salt water

salt water

one clear liquid: uniform throughout, every drop the same

Ocean water is salt dissolved in water.

A common first attempt: uniform throughout, so a compound. Test it.

✓ Worked example 2: solution

salt water

one clear liquid: uniform throughout, every drop the same

A common first attempt

salt water = a compound?

a compound keeps one fixed ratio; this sample holds however much salt was stirred in ✕

Step 1 · Count the kinds of particles

Step 2 · Count the elements or the phases

Step 3 · Name the class

salt water → a homogeneous mixture

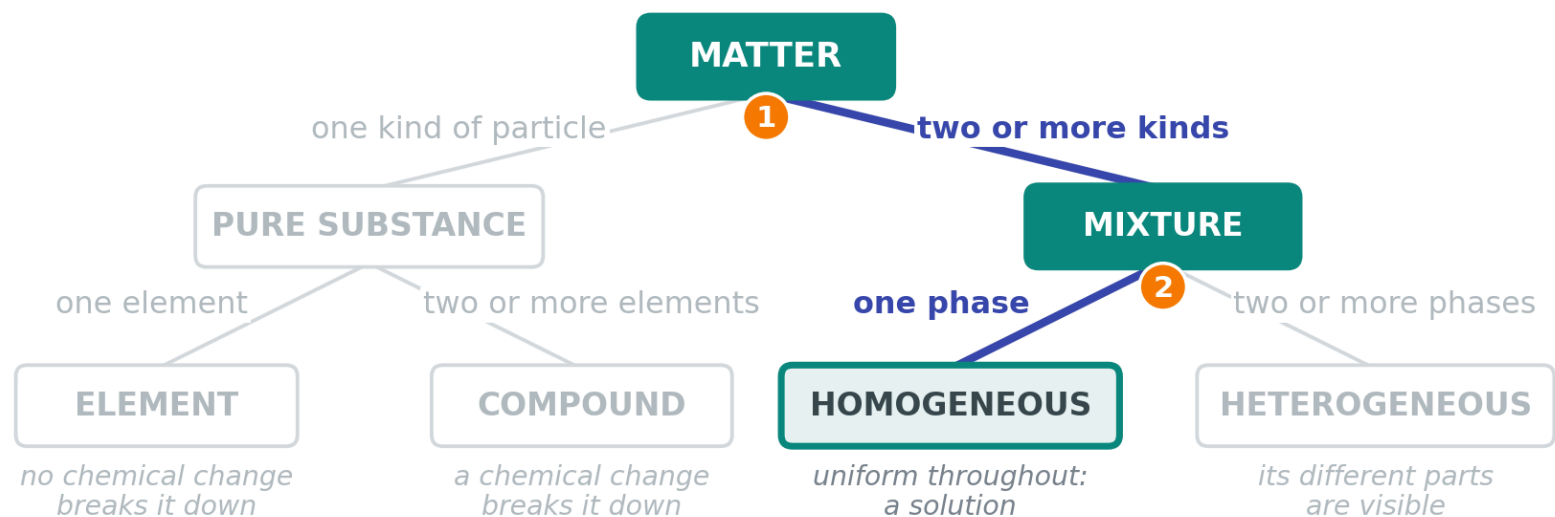
two kinds of particles · one phase · composition varies

 **Sense check:** Uniform answers the phase question, not the purity question. Salt water is uniform and still a mixture.

✓ Worked example 2: the route on the map

salt water

given: water and dissolved salt · found: a homogeneous mixture



🧠 **Sense check:** Count 1: two kinds of particles → mixture. Count 2: one phase → homogeneous. The route never visits the compound box. ✓

Take-home: uniform does not mean compound

water: always 2 H : 1 O

compound: one substance, fixed ratio, separated only by chemical reaction

salt water: any ratio that dissolves

homogeneous mixture: uniform, variable composition, separated by boiling

A compound keeps one fixed ratio. A homogeneous mixture is uniform, but its composition can vary. Uniformity describes phases, never bonding.

Your turn: oil-and-vinegar dressing

oil-and-vinegar dressing
an oil layer floating on a vinegar layer

step	question	answer
1 · kinds of particles	one kind, or more?	more than one → _____
2 · phases	how many phases?	_____ distinct layers
3 · name the class		_____

Complete the three counts.

dressing → a heterogeneous mixture
more than one kind of particle · two phases: a top sample is oil, a bottom sample is vinegar

⚠ Where this goes wrong

Calling a uniform mixture a compound. Brass is uniform, but its copper-to-zinc ratio varies batch to batch. A compound keeps one fixed ratio. Uniform appearance does not show chemical bonds.

Calling every multi-substance sample heterogeneous. Air holds nitrogen, oxygen, and argon in one phase. Several substances can share a single uniform phase: homogeneous.

Reading "same properties throughout" as pure. Same everywhere means one phase, nothing more. A pure substance also needs fixed composition, and sugar water's composition can vary.



Practice 1

white vinegar: acetic acid dissolved in water
one clear liquid, uniform throughout

A bottle of white vinegar looks completely uniform. How should it be classified, and why?

- A. Pure substance: it shows the same properties at every point, so it is a single substance
- B. Homogeneous mixture: it is uniform throughout, but its acid-to-water ratio can vary
- C. Compound: a uniform liquid must have its components chemically bonded in a fixed ratio
- D. Heterogeneous mixture: it contains more than one substance, so it cannot be uniform

✓ Practice 1 · answer: B

white vinegar → a homogeneous mixture (answer B)

two kinds of particles · one phase · one bottle can hold more acid than another

A: same properties throughout shows one phase, not one substance; the composition can still vary. C: uniform appearance does not show bonding; the acid and water separate by distillation, a physical change. D: several substances can share one phase; classification follows the phase count, not the substance count.

 **Sense check:** Uniform → homogeneous. Variable composition → mixture. Both labels apply to the same bottle.



Worked example 3: air, brass, milk

air · brass · milk

a gas, a solid, and a liquid

The air in the room, the brass of a doorknob, a glass of milk. Work the three steps on each sample.

✓ Worked example 3: kinds of particles

air · brass · milk
a gas, a solid, and a liquid

Step 1 · Count the kinds of particles

sample	particles
air	N ₂ , O ₂ , Ar, and more
brass	copper atoms and zinc atoms
milk	water, fats, proteins, sugars

Each sample holds more than one kind of particle: three mixtures, in three physical states.

 **Sense check:** A mixture can be a gas, a solid, or a liquid. State does not enter the classification.

✓ Worked example 3: phases and the class

air · brass · milk: three mixtures

Step 2 · Count the elements or the phases

sample	phases
air	one, uniform at every point
brass	one, a uniform solid
milk	two, fat droplets in a watery liquid

Step 3 · Name the class

air → homogeneous · brass → homogeneous · milk → heterogeneous

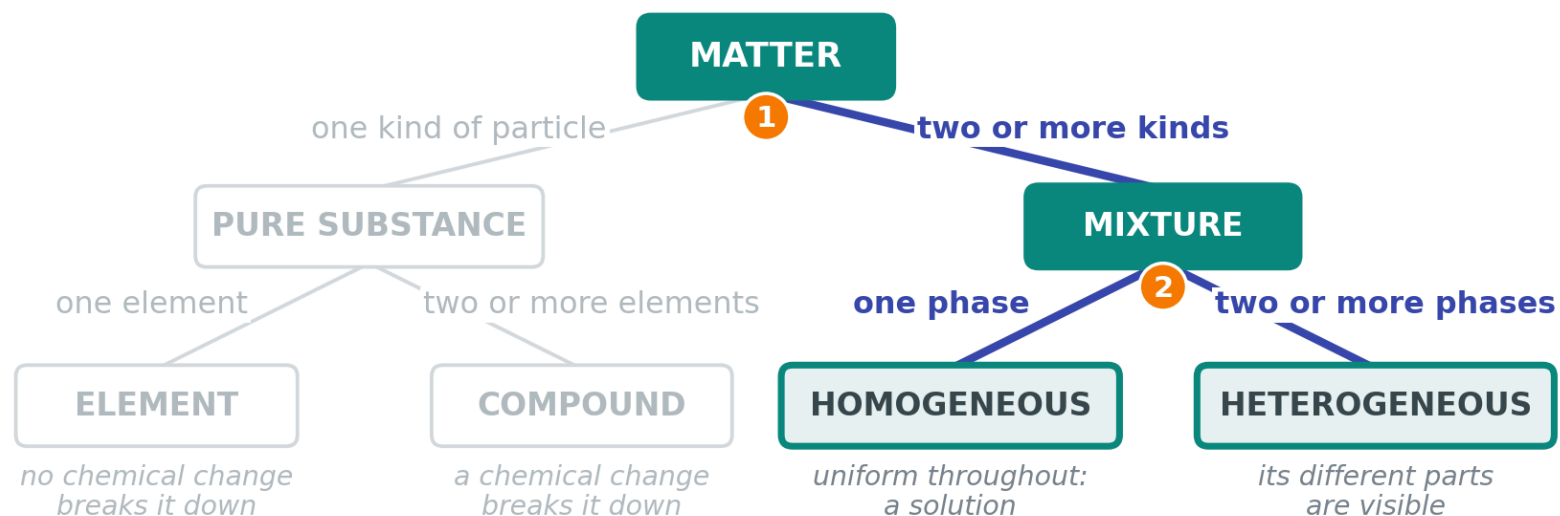
air: $78\% \text{ N}_2 + 21\% \text{ O}_2 + 1\% \text{ other} = 100\%$ · one phase, variable composition

 **Sense check:** Milk looks uniform; magnified, fat droplets show real boundaries. Phase count, not appearance.

✓ Worked example 3: the route on the map

air · brass · milk

given: more than one kind of particle in each · found: air, brass homogeneous · milk heterogeneous



① count the kinds of particles ② count the elements (pure) or the phases (mixture) ③ name the class

Sense check: Count 1 sends all three to the mixture side. Count 2 splits them: one phase for air and brass, two for milk. ✓



Practice 2

an unlabeled jar of colorless crystals

observation 1: under magnification, every crystal looks identical · observation 2: every sample analyzed has the same composition; strong heating leaves black carbon as water vapor escapes

How should the crystals be classified?

- A. Homogeneous mixture: every crystal looks alike, and heating separates the carbon from the water
- B. Element: every crystal is identical, and every sample has the same composition
- C. Compound: fixed composition in every sample, and heating breaks it into new, simpler substances
- D. Heterogeneous mixture: heating leaves a black solid and a vapor, two phases

✓ Practice 2 · answer: C

the crystals → a compound (answer C)

one kind of particle · fixed composition · a chemical change breaks it down

Step 1: the same composition in every sample means one kind of particle, a pure substance. Step 2: heating turns it into carbon and water, new substances: a chemical change. A pure substance that a reaction breaks down holds more than one element: a compound.

A: a mixture varies in composition and separates by a physical change; charring is chemical. B: fixed composition marks any pure substance, and no reaction breaks an element down. D: counts the phases of the products, not of the sample.

 **Sense check:** Uniform describes phases. Fixed composition says pure. Breakdown by reaction says compound.



Check yourself

1. A sealed bottle of soda water looks uniform throughout. Work the counts: kinds of particles, then phases. What class results?
2. Ice floats in liquid water. How many kinds of particles? How many phases? Is the sample a mixture?

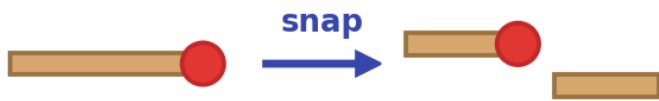
Separating a mixture is a **physical change**: boiling, filtering, settling. Breaking a compound into its elements is a **chemical change**. The same distinction, physical or chemical, classifies every property a substance shows.

3 · Physical & Chemical Properties

Classify any property or change as physical or chemical by asking whether the substance keeps its identity, read the four signals of a chemical change, and sort intensive from extensive properties with the halving test.

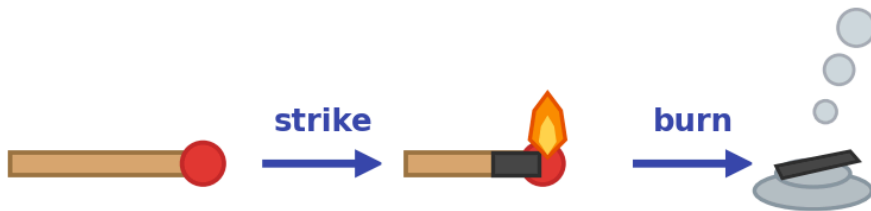
One match, two fates

which change kept the wood?



still wood ✓

two pieces, same substance



wood gone ✗

flame gases + ash: new substances

no cooling, sorting, or gluing brings the wood back

Snap a match: two pieces, still wood. Strike it: flame, smoke, ash, and nothing brings the wood back. Some changes keep a substance. Some end it.

Two kinds of properties

A property describes a substance; the question is what the description costs. Physical properties can be observed with the sample intact. Chemical properties tell what a substance can become, and seeing that takes a reaction.

physical property: observe it, keep the sample

color · odor · density · melting point: the coin survives having its density measured

chemical property: see it only in a reaction

flammable · rusts · reacts with acid: the match that proves flammability is spent proving it

Two kinds of changes

property or change · physical or chemical

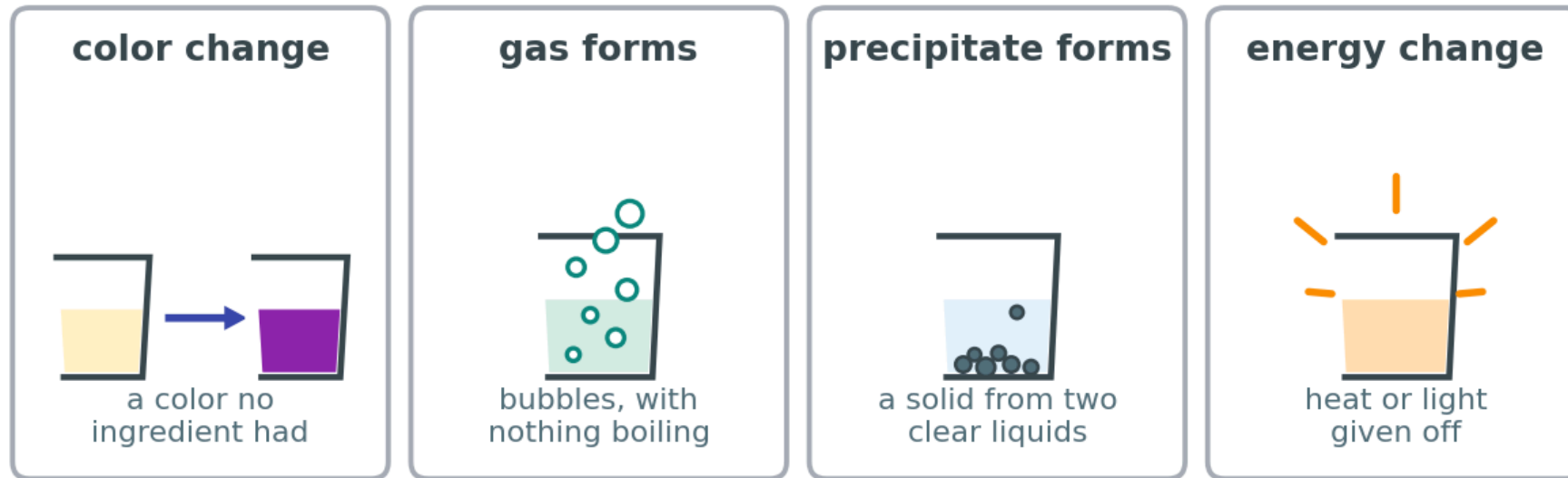
	physical identity kept	chemical identity changed
property a description	melts at 35 °C · 2.70 g/mL measure it, keep the sample	flammable · can rust seen only in a reaction
change an event	ice melts · sugar dissolves same substance, new form	toast browns · iron rusts new substance formed

properties describe, changes happen; both sort by one question: did the substance survive?

Changes get the identity question too. Melt ice, dissolve sugar, tear paper: the substances are all still there, in new forms. Burn paper and it is gone; new substances hold its atoms.

Evidence a new substance formed

four signals that a new substance formed



strong hints, not proof: the bubbles in boiling water are still just water

No one watches atoms rebond. What shows up instead: a color no ingredient had, a gas from a mixture that is not boiling, a solid from two clear liquids, heat or light given off.

Intensive or extensive

Some physical properties measure how much sample sits on the bench. Others describe the substance itself, at any amount. Only the second kind identifies an unknown; a pile's mass says nothing about what the pile is.

extensive: depends on the amount

mass · volume · length: halve the sample and the value halves · EXtensive: follows the EXtent

intensive: ignores the amount

density · temperature · melting point: halve the sample and the value holds · INtensive: INdependent of amount



The method

1. **Name the substance before and after.** Same substance → physical. New substance → chemical.
2. **Check the evidence.** New color, gas, solid, or energy change → new substance.
3. **Halve the sample mentally.** Value halves → extensive. Value holds → intensive.



Worked example 1: butter in a pan

butter in a hot pan: melts, then browns and smokes

event 1: solid → clear yellow liquid · event 2: liquid → brown residue + smoke + sharp smell

Butter dropped in a warm pan melts. Left on the heat, it browns, smokes, and turns bitter. Two events, one pan. Classify each.

✓ Worked example 1: solution

butter in a hot pan: melts, then browns and smokes

Step 1 · Name the substance before and after

melting → a physical change

same butter · new form · fully undone by cooling

Step 1 · Name the substance before and after, again

Step 2 · Check the evidence

browning → a chemical change

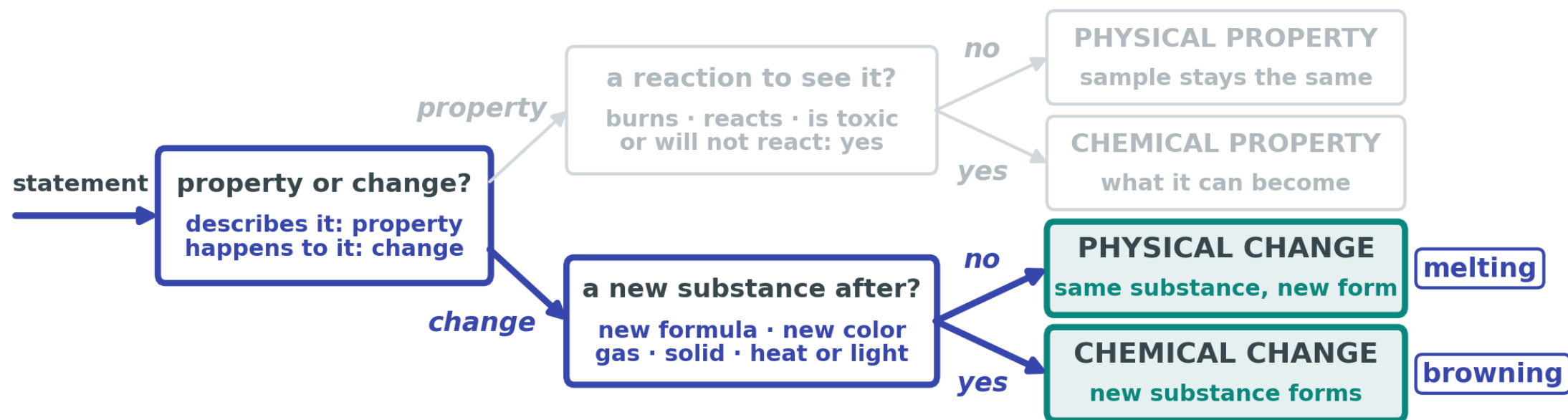
color change · gas and smoke · new smell · not undone by cooling

 **Sense check:** The pan ran the identity test twice. Cooling reverses melting; nothing on a stove has ever unbrowned toast, and butter follows the same rule.

✓ Worked example 1: the route on the map

butter in a hot pan: melts, then browns and smokes

melting: same butter → physical change · browning: brown solids, smoke, new smell → chemical change



🧠 **Sense check:** One pan, two exits. Both events are changes; only browning made a new substance. ✓



Worked example 2: sawing a metal bar

a shiny gray bar: mass 10.0 g · volume 3.70 mL

sawed into two equal pieces · wanted: which numbers change, which still identify the metal

A bar of unknown metal is sawed in half. A common first claim: cutting halves the mass, so every property must shrink with it. Test each number.

✓ Worked example 2: solution

Step 1 · Name the substance before and after

Step 2 · Check the evidence

Step 3 · Halve the sample mentally

whole bar: $10.0 \text{ g} \div 3.70 \text{ mL} = 2.70 \text{ g/mL}$

half bar: $5.00 \text{ g} \div 1.85 \text{ mL} = 2.70 \text{ g/mL}$

mass and volume halved → extensive · density held → intensive

$10.0 \rightarrow 5.00 \text{ g} \cdot 3.70 \rightarrow 1.85 \text{ mL} \cdot 2.70 \text{ g/mL}$ before and after

 **Sense check:** The saw changed how much metal sits on the bench, never what the metal is. The identity number did not flinch.

Take-home: one question does the sorting

identity kept → physical

melting, dissolving, cutting: the substance survives in a new form

identity lost → chemical

burning, rusting, browning: new substances hold the old atoms

Every label in this topic rides on one question: is the original substance still there? Yes, in any shape or state: physical. No: chemical, and the evidence usually announces it.

Your turn: a nail in the rain

an iron nail, weeks outdoors: an orange-brown, flaky coat
the coat scrapes off as a brittle powder, nothing like the shiny metal beneath

step	question	answer
1 · substance after	is the coat still iron?	no, a new orange solid → _____ change
2 · evidence	which signals appear?	a _____ change and a new solid
3 · undo it	does scraping restore the iron?	_____

Complete the three rows.

rusting → a chemical change

iron + air + water → rust · color change, new brittle solid · scraping removes rust and restores nothing

⚠ Where this goes wrong

Calling dissolving a chemical change. Sugar vanishes into tea, but boil the tea away and the sugar is back, unchanged. Out of sight is not out of existence: the sugar kept its identity. Physical.

Reading every bubble as a reaction. Boiling water bubbles furiously, and the gas is water vapor: the same substance. Gas counts as evidence when it is a new substance appearing from a mixture that is not boiling.

Confusing a property with a change. "Melts at 35 °C" is a physical property, a number that sits on a label. "The butter melted" is a physical change, an event in a pan. Properties describe; changes happen.

Treating density as extensive. Twice the metal weighs twice as much, so the density gets doubled too. But mass and volume both doubled, and their ratio went nowhere. Density ignores amount; that is why it identifies.



Practice 1

gasoline: a liquid · floats on water · evaporates fast · flammable
four descriptions of one substance

Four descriptions of gasoline. Which one is a chemical property?

- A. It is a liquid at room temperature
- B. It floats on water
- C. It evaporates quickly from an open container
- D. It is flammable

✓ Practice 1 · answer: D

flammable → a chemical property (answer D)

observing it turns the gasoline into carbon dioxide and water: the description costs the sample

A: state is observed by looking; the liquid stays gasoline. B: floating compares two densities, and both liquids survive the comparison. C: evaporation trades liquid for vapor, still gasoline; chill the vapor and it condenses right back.

 **Sense check:** Three descriptions leave the gasoline in the can. The fourth can be confirmed once, from a distance.



Practice 2

a copper wire, four entries in a lab report

(1) mass: 45.0 g · (2) density: 8.96 g/cm³ · (3) forms a green coating after years in moist air · (4) melts at 1085 °C

Which classification of the four entries is correct?

- A. (3) and (4) are chemical properties: observing either one ruins the wire
- B. (2) and (4) are intensive physical properties that identify copper; (3) is a chemical property; (1) is extensive
- C. (1) and (2) are extensive: a longer wire has more mass and more density
- D. (3) and (4) are changes, not properties; only (1) and (2) describe the copper

✓ Practice 2 · answer: B

(1) extensive · (2), (4) intensive, physical · (3) chemical (answer B)

halve the wire: $45.0 \div 2 = 22.5 \text{ g}$ · still 8.96 g/cm^3 · still melts at $1085 \text{ }^\circ\text{C}$

Melted copper is still copper, so (4) is physical; the green coating is a new substance, so (3) is chemical. The halving test sorts the numbers: mass halves, density and melting point hold.

A: melting changes the form, never the identity. C: density held when the wire was halved; mass and volume shrink together. D: "melts at $1085 \text{ }^\circ\text{C}$ " is a number on a label, not an event on the bench. Properties describe; changes happen.

 **Sense check:** Two questions per entry: what survived, and does halving change the number. Only intensive physical properties identify a sample.

Check yourself

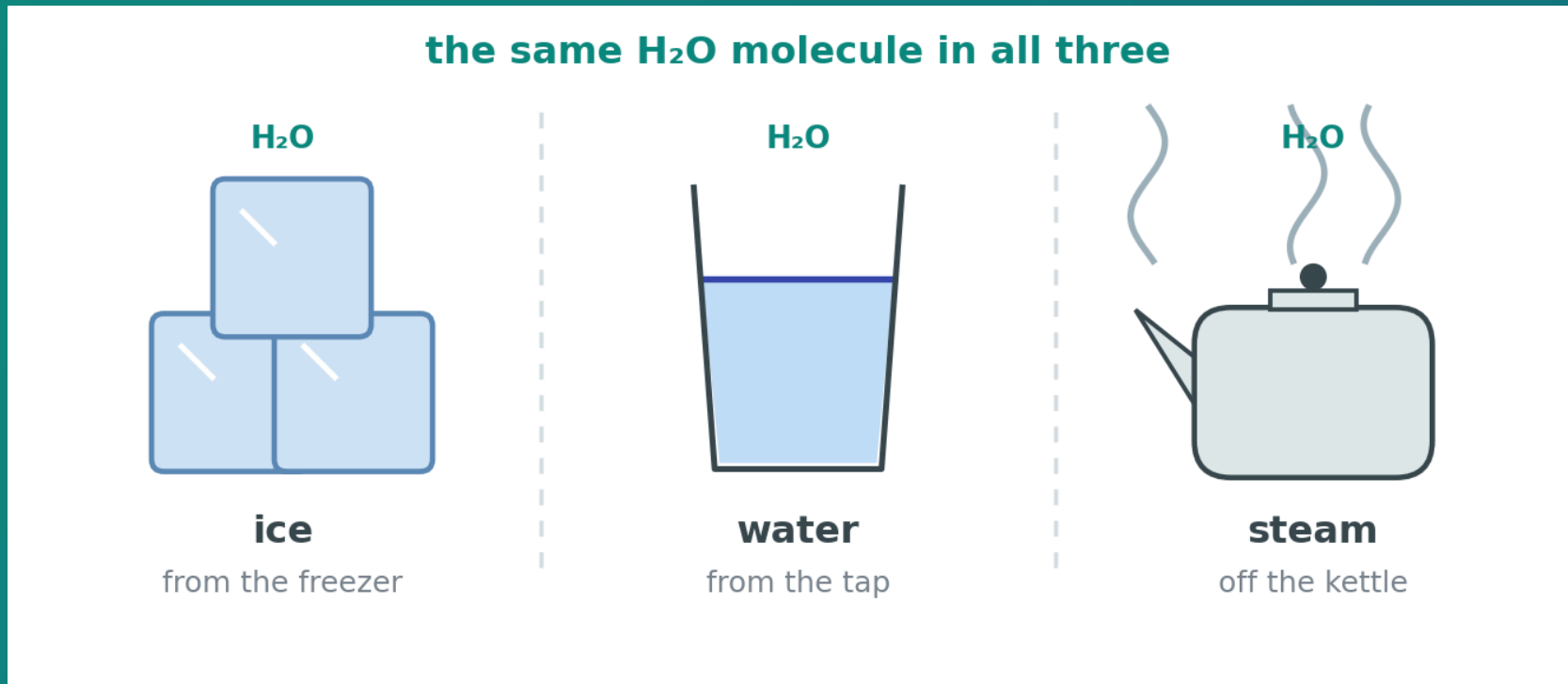
1. A tablet dropped in water fizzes, the glass cools, and the tablet shrinks away. List every signal you can, then classify the change.
2. An unknown liquid arrives with four measurements: mass, volume, density, boiling point. Which two could identify it, and why do the other two fail?

Physical or chemical, property or change, intensive or extensive: every label here comes from one habit. Ask what survived the event, and what the observation cost.

4 · States of Matter

Name a sample's state from two checks, its shape and its volume, and back the call with the particle picture of spacing and motion.

One substance, three states

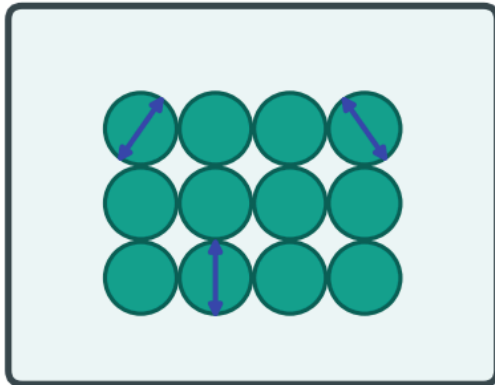


Ice, tap water, steam off the kettle: all H₂O, the same molecule down to the last atom. What changed is how the molecules sit and how fast they move.

The particle view

same particles, three arrangements

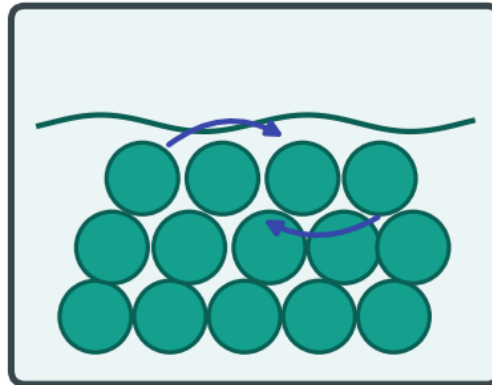
solid



particles vibrate in place

**own shape
own volume**

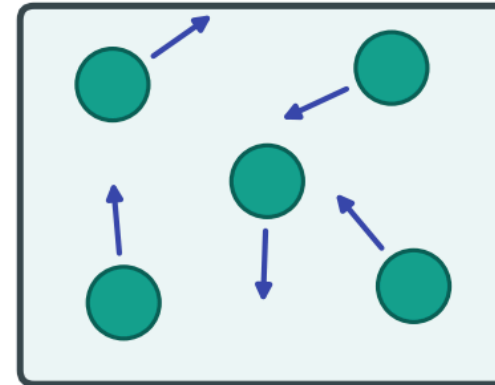
liquid



particles slide past one another

**container's shape
own volume**

gas



particles fly between collisions

**container's shape
container's volume**

spacing and motion name the state

Matter is particles in motion. Solid: the particles touch and vibrate in place. Liquid: they touch but slide past one another. Gas: they fly apart, mostly empty space between them.

Shape and volume follow the particles

solid: its own shape, its own volume
locked particles hold both

liquid: the container's shape, its own volume
touching holds the volume · sliding gives the shape away

gas: the container's shape, the container's volume
far-apart particles spread until the container stops them

Two questions sort every sample: does it keep its own shape, and does it keep its own volume. The particle picture answers both.

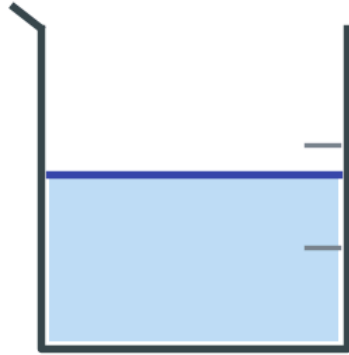
The three states in glassware

three states on the bench



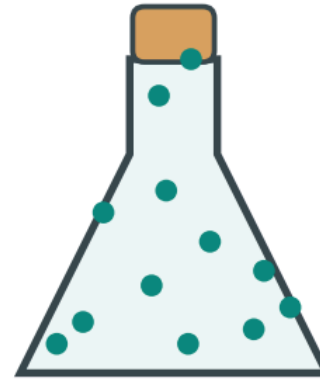
solid

the powder keeps its heap



liquid

levels flat, takes the beaker's shape



gas

fills the flask, neck included

The bench shows all three daily. A powder holds its heap on the watch glass. A liquid levels flat against the beaker walls. A gas fills a stoppered flask completely, neck included.



Practice

a few drops of liquid bromine in a stoppered flask
minutes later: brown vapor fills the whole flask, neck included

A student explains: "As a gas, the bromine particles swell to many times their size, so they fill the flask." What is the flaw?

- A. Nothing: gas particles are larger than liquid particles, so a gas takes up more room
- B. Size is not the reason; the vapor's particles still touch, and sliding carries them into every corner
- C. The particles keep their size; as a gas they fly apart, with mostly empty space between them
- D. Size is not the reason; evaporating made a new, lighter substance that rises to fill the flask

✓ Practice · answer: C

same particles, new spacing (answer C)

liquid: touching, its own volume · gas: far apart, the container's volume

A accepts the claim; particles keep their size in every state, and the gaps between them grew. B describes a liquid; touching particles hold their own volume, and the vapor took the flask's. D breaks the identity rule; the vapor is the same bromine, molecule for molecule, and cooling returns the liquid.

 **Sense check:** A state change moves the particles apart or together. It never resizes them and never makes a new substance.



Check yourself

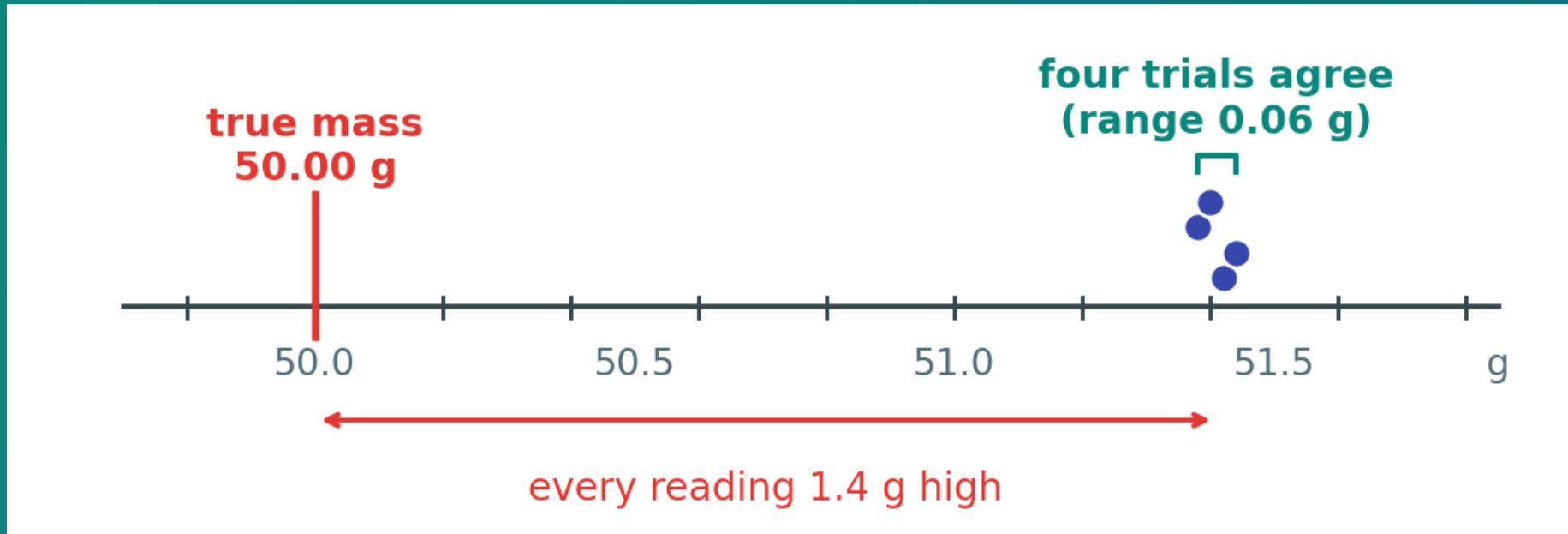
1. A sealed syringe of air squeezes to half its size; a sealed syringe of water barely gives at all. Explain both results with particle spacing.
2. Steam from a shower fogs the mirror into liquid droplets. Did the molecules themselves change, or only their spacing and motion?

Every change of state carries an energy cost: melting and boiling take energy in, freezing and condensing give it back. Measuring that energy is its own topic, and it starts from this particle picture.

5 · Precision & Accuracy

Judge a set of repeated measurements two ways: precision from the trials' relative range, accuracy from their average's percent error, each verdict decided against the 2% bar.

The same wrong answer, four times



A 50.00-g standard is weighed four times. Every reading lands near 51.4 g. The trials agree with one another; not one of them is right.

Two questions about repeated measurements

Repeated trials form a set, and the set is judged twice. Each verdict comes from its own comparison.

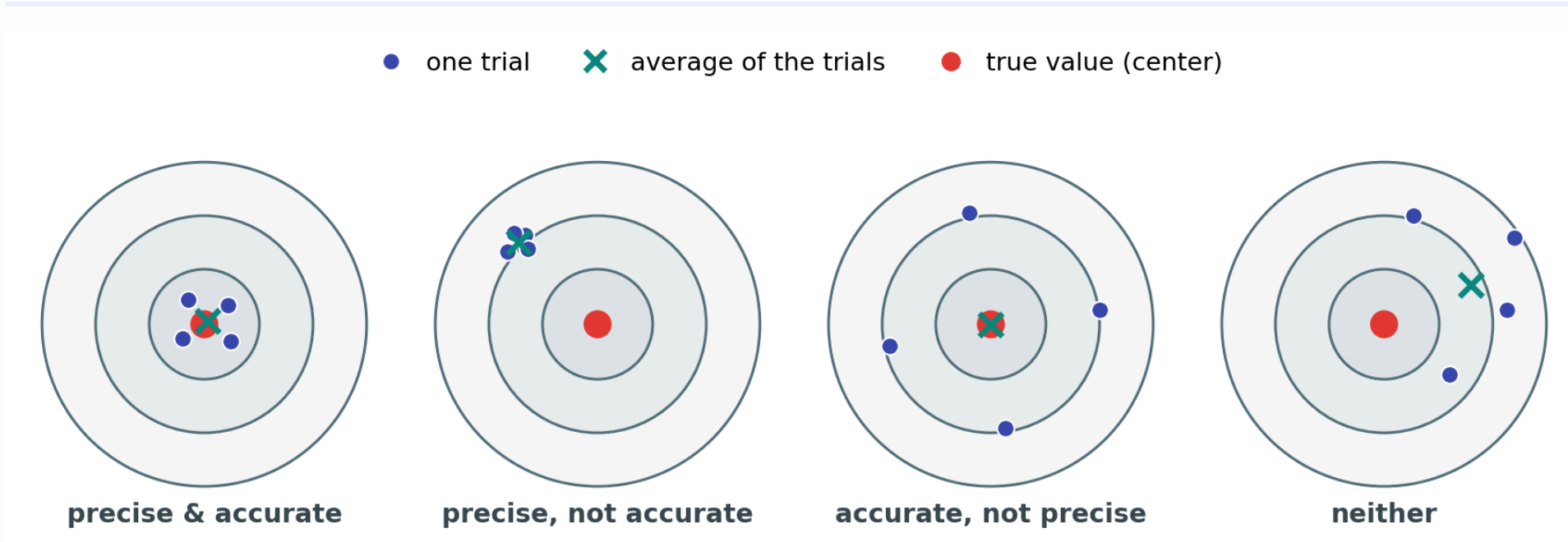
precise = the trials agree with one another

uses only the trials themselves · pRecise: Repeatable

accurate = the average lands on the true value

uses the average and the accepted true value · aCcurate: Correct

The four outcomes



The center is the true value; × marks the average. A shared flaw, a miscalibrated balance, shifts the whole cluster together: precise but wrong. Random scatter cancels in the average: accurate but imprecise. The verdicts are independent.

A number for each verdict

Each verdict has its own number. The range measures agreement among the trials. The error measures how far the average sits from the true value.

range = highest trial – lowest trial

$\text{range} \div \text{average} \times 100 = \text{relative range (precision's number)}$

error = average – true value

$\text{error} \div \text{true value} \times 100 = \text{percent error (accuracy's number)}$

The 2% standard

"Close" and "consistent" are not measurements. Compute both percentages; test each against the same bar: 2%.

relative range = range ÷ average × 100; at most 2% → precise
above 2%, the trials disagree: not precise

percent error = |average – true value| ÷ true value × 100; at most 2% → accurate
above 2%, the average misses: not accurate

The 2% bar is a course convention, not a law of nature; other labs draw the line elsewhere.



The method

1. **Compute the range**: highest trial – lowest.
2. **Compute the average** of the trials.
3. **Judge precision**: relative range at most 2%.
4. **Judge accuracy**: percent error at most 2%. State both verdicts.



Worked example 1: checking a balance

trials: 51.42 g · 51.38 g · 51.44 g · 51.40 g

true value: 50.00 g (calibration standard) · wanted: both verdicts

A balance is checked with a standard of known mass, weighed four times.

Judge the set: precise? accurate?

✓ Worked example 1: solution

Step 1 • Compute the range

Step 2 • Compute the average

range: $51.44 \text{ g} - 51.38 \text{ g} = \mathbf{0.06 \text{ g}}$

average: $(51.42 + 51.38 + 51.44 + 51.40) \div 4 = 205.64 \div 4 = \mathbf{51.41 \text{ g}}$

Step 3 • Judge precision

Step 4 • Judge accuracy

relative range: $0.06 \div 51.41 \times 100 = \mathbf{0.12\%} \rightarrow$ at most 2%: precise

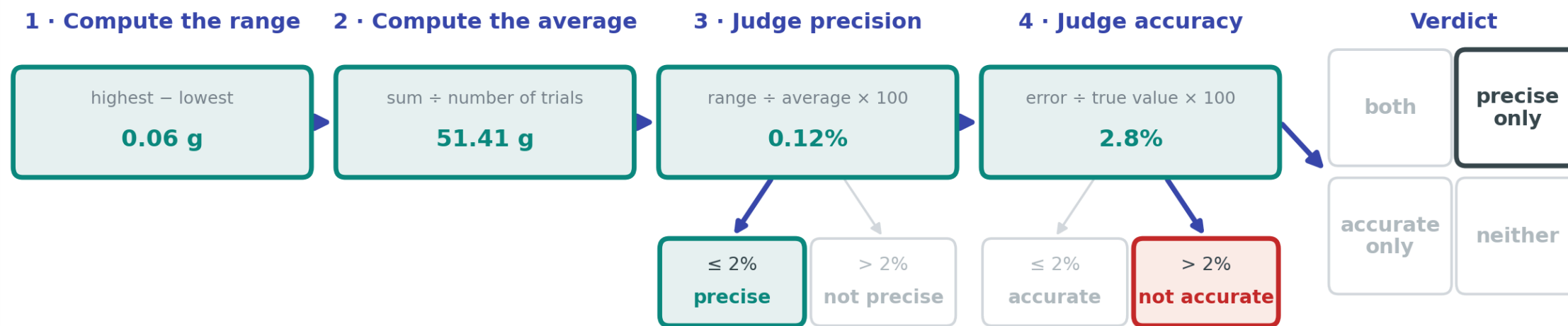
error: $51.41 \text{ g} - 50.00 \text{ g} = \mathbf{1.41 \text{ g high}}$

$\rightarrow$ percent error $1.41 \div 50.00 \times 100 = \mathbf{2.8\%}$, above 2%: not accurate

 **Sense check:** Precise but not accurate. All four trials carry the same 1.41-g error; agreement cannot expose it.

✓ Worked example 1: the route on the map

trials: 51.42 · 51.38 · 51.44 · 51.40 g · true value: 50.00 g
 found: range 0.06 g · average 51.41 g · relative range 0.12% · percent error 2.8%



🧠 **Sense check:** The forks split: under 2% at precision, above 2% at accuracy. The set lands in "precise only". ✓



Worked example 2: boiling water

trials: $97.9\text{ }^{\circ}\text{C}$ · $102.1\text{ }^{\circ}\text{C}$ · $99.6\text{ }^{\circ}\text{C}$ · $100.4\text{ }^{\circ}\text{C}$

true value: $100.0\text{ }^{\circ}\text{C}$ (water boils at sea level) · wanted: both verdicts

A thermometer is read four times in boiling water at sea level.

A common first answer: the readings all land near the true $100.0\text{ }^{\circ}\text{C}$, so they are precise.
Test it against the 2% standard.

✓ Worked example 2: solution

Step 1 · Compute the range

Step 2 · Compute the average

range: $102.1\text{ }^{\circ}\text{C} - 97.9\text{ }^{\circ}\text{C} = 4.2\text{ }^{\circ}\text{C}$

average: $(97.9 + 102.1 + 99.6 + 100.4) \div 4 = 400.0 \div 4 = 100.0\text{ }^{\circ}\text{C}$

Step 3 · Judge precision

Step 4 · Judge accuracy

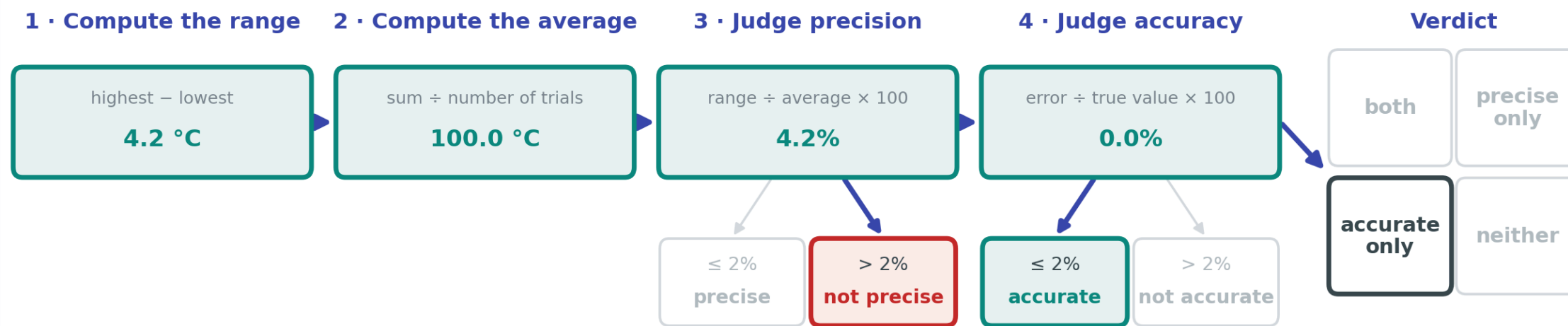
relative range: $4.2 \div 100.0 \times 100 = 4.2\%$ → above 2%: not precise

percent error: $(100.0 - 100.0) \div 100.0 \times 100 = 0.0\%$ → at most 2%: accurate

 **Sense check:** Accurate but not precise. Random errors land high and low with equal chance; in the average they cancel.

✓ Worked example 2: the route on the map

trials: 97.9 · 102.1 · 99.6 · 100.4 °C · true value: 100.0 °C
 found: range 4.2 °C · average 100.0 °C · relative range 4.2% · percent error 0.0%



🧠 **Sense check:** The opposite split: above 2% at precision, under 2% at accuracy. The set lands in "accurate only". ✓

Take-home: two separate comparisons

trials: 97.9 · 102.1 · 99.6 · 100.4 °C; true value 100.0 °C

relative range 4.2%, above the 2% bar → not precise · percent error 0.0%, under it → accurate

Precision compares the trials with one another and never mentions the true value. Accuracy compares the average with the true value and never mentions the spread. Swapping the two comparisons flips the verdict.

Your turn: density of an aluminum sample

trials: 2.41 · 2.43 · 2.40 · 2.42 g/mL

true density of aluminum: 2.70 g/mL

range: $2.43 - 2.40 = \underline{\hspace{1cm}}$ g/mL · average: $(2.41 + 2.43 + 2.40 + 2.42) \div 4 = \underline{\hspace{1cm}}$ g/mL

range: $2.43 - 2.40 = \mathbf{0.03 \text{ g/mL}}$

→ relative range $0.03 \div 2.415 \times 100 = \mathbf{1.2\%}$, at most 2%: precise

average: $9.66 \div 4 = \mathbf{2.415 \text{ g/mL}}$

→ percent error $0.285 \div 2.70 \times 100 = \mathbf{10.6\%}$, above 2%: not accurate

Precise but not accurate.

⚠ Where this goes wrong

Swapping the definitions. Trials of 97.9–102.1 °C averaging 100.0 °C get called "precise, since the average is right." An average on the true value is accuracy. Precision is agreement among the trials: relative range 4.2%, above 2% → not precise.

Judging the set by one trial. In 51.42, 51.38, 51.44, 51.40 g against a true 50.00 g, trial 2 sits 51.38 – 50.00 = 1.38 g from the true mass, yet the set is precise: relative range 0.12%, well under 2%. Both verdicts describe the whole set, never one reading.

Counting digits as accuracy. A caliper reads 14.42 cm, four digits, on a rod whose true length is 15.00 cm: the reading is 0.58 cm short. Digits show how finely the scale reads, not whether the reading is right.

Treating precise as accurate. Four readings within 0.06 g of one another all sit 1.41 g above the true mass: percent error 2.8%, above the bar. Agreement rules out scatter; it cannot rule out an error shared by every trial.



Practice 1

trials: 26.32 mL · 26.28 mL · 26.34 mL · 26.30 mL

true volume: 25.00 mL

A pipette made to deliver 25.00 mL is tested four times. Which statement gives both verdicts?

- A. Precise but not accurate: the trials agree within 0.06 mL, and their average of 26.31 mL is 1.31 mL above the true volume.
- B. Not precise: trial 2, at 26.28 mL, sits 1.28 mL from the true volume, and one reading that far off rules out precision.
- C. Accurate but not precise: agreement within 0.06 mL makes the trials accurate, and the 1.31-mL gap from the true volume makes them imprecise.
- D. Accurate: every reading carries four digits, and digits that fine mean accuracy.

✓ Practice 1 · answer: A

trials: 26.32 mL · 26.28 mL · 26.34 mL · 26.30 mL

true volume: 25.00 mL

range: $26.34 - 26.28 = 0.06 \text{ mL}$

→ relative range $0.06 \div 26.31 \times 100 = 0.23\%$, at most 2%: precise

average: $(26.32 + 26.28 + 26.34 + 26.30) \div 4 = 26.31 \text{ mL}$

→ percent error $1.31 \div 25.00 \times 100 = 5.2\%$, above 2%: not accurate → **answer A**

 **Sense check:** All four deliveries run high by nearly the same amount. A flaw in the pipette itself repeats identically in every trial. ✓



Practice 2

readings: 4.91 g · 5.09 g · 4.98 g · 5.06 g
true mass: 5.00 g

A 5.00-g brass check weight goes on a pocket scale four times. Which verdict is correct?

- A. Precise and accurate: the readings spread only 0.18 g and the average misses by only 0.01 g, both well under 2
- B. Precise but not accurate: the 0.2% figure shows the readings agree, and the 3.6% figure shows the average misses
- C. Neither: a 3.6% spread fails the bar, and readings that scatter cannot average onto the true value
- D. Accurate but not precise: the relative range, 3.6%, fails the 2% bar; the percent error, 0.2%, passes it

✓ Practice 2 · answer: D

range: $5.09 - 4.91 = 0.18 \text{ g}$

→ relative range $0.18 \div 5.01 \times 100 = 3.6\%$, above 2%: not precise

average: $20.04 \div 4 = 5.01 \text{ g}$

→ percent error $0.01 \div 5.00 \times 100 = 0.2\%$, at most 2%: accurate → **answer D**

A treated the 2% bar as 2 grams: 0.18 g is small only next to a large mass, and on a 5-g weight it is 3.6%. B swapped the definitions: relative range judges precision, percent error judges accuracy. C let one verdict answer both: high and low readings cancel in the average.

🧠 **Sense check:** A small mass makes a small spread large in percent. Divide before judging. ✓

Check yourself

1. Four trials of a 20.00-g standard have a range of 0.04 g, and their average, 17.90 g, sits 2.1 g below the true value. Compute both percentages and give both verdicts.
2. A single reading lands exactly on the true value. What does that establish about the set's precision?

Every trial in these sets was recorded to a fixed decimal place. How many digits a measurement may claim is its own rule set: **significant figures**.

6 · Scientific Notation

Write numbers in scientific notation and read them back (the decimal's move sets the exponent's size and sign, the coefficient stays at $1 \leq M < 10$), and enter them on a calculator with EE.

Numbers too big and too small to write out

Two numbers chemistry runs on

particles in one mole

602,214,076,000,000,000,000,000

$$= 6.022 \times 10^{23}$$

15 zeros just to hold the size

mass of one hydrogen atom

0.000000000000000000000000167 g

$$= 1.67 \times 10^{-24} \text{ g}$$

23 zeros just to hold the size

the coefficient keeps the digits · the exponent carries the size

Written out, chemistry's numbers are unreadable. Scientific notation writes each in a handful of characters: the digits once, then a power of ten carrying the size.

One coefficient, one exponent

Scientific notation writes any number as a coefficient times a power of ten: the coefficient carries the digits, the exponent carries the size.

$$6.022 \times 10^{23}$$

coefficient M with $1 \leq M < 10$ · exponent n is a whole number · $10^{23} = \text{"times ten, 23 times"}$

Exactly one nonzero digit stands before the decimal point.

Standard → scientific: count the decimal's move

Slide the decimal until one nonzero digit leads. Each factor of ten moves it one place, so places moved = factors of ten = the exponent n . Direction sets the sign.

Moving the decimal point sets the exponent

how far it moves = the size of n · which way it moves = the sign of n

move 3 places right →

0 . 0 0 4 5 6



$$= 4.56 \times 10^{-3}$$

number < 1 → negative exponent

move 7 places left ←

9 3 0 0 0 0 0 0



$$= 9.3 \times 10^7$$

number > 1 → positive exponent

Write it as $M \times 10^n$ with $1 \leq M < 10$

moved RIGHT (was < 1) → $n < 0$ · moved LEFT (was > 1) → $n > 0$



The method: convert to scientific notation

1. Place the decimal so one nonzero digit leads: that is M .
2. Count the places moved: that is $|n|$.
3. Sign: moved left $\rightarrow$ n positive; moved right $\rightarrow$ negative.
4. Write $M \times 10^n$.



Worked example 1: two conversions

0.00456 and 93,000,000

write each in scientific notation, $M \times 10^n$

One is small (less than 1), one is large (greater than 1). Count the decimal's move for each, then read the sign from the direction.

✓ Worked example 1: solution

The small number moves right (negative n); the large number moves left (positive n).

$$0.00456 = 4.56 \times 10^{-3}$$

$$93,000,000 = 9.3 \times 10^7$$

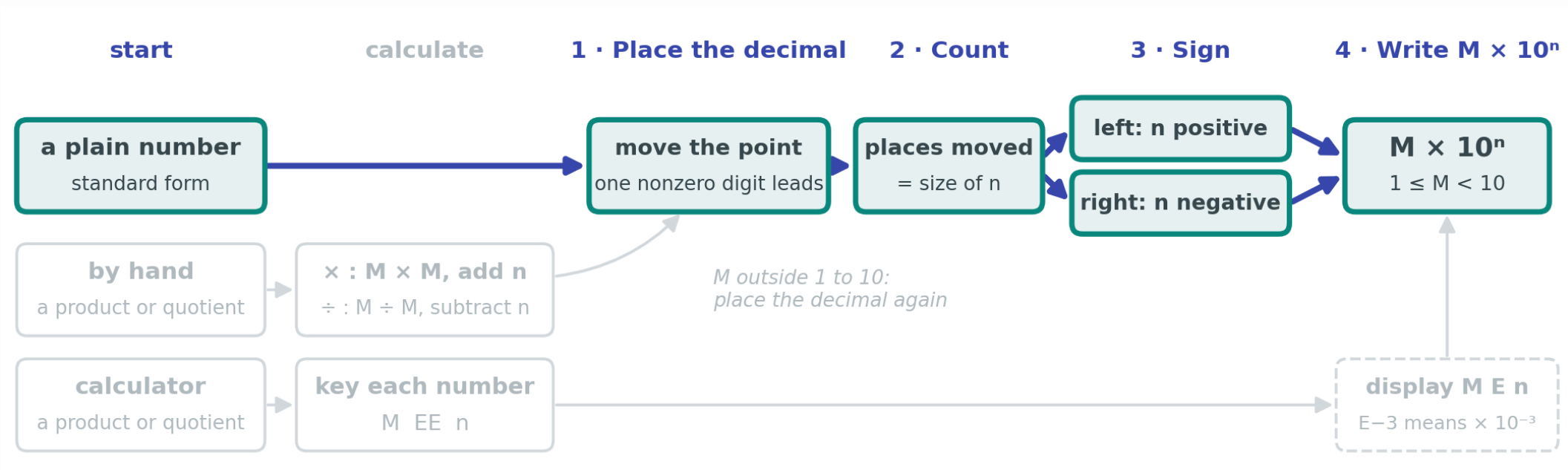
3 places right $\rightarrow n = -3$ · 7 places left $\rightarrow n = +7$

🧠 **Sense check:** A leading zero forces n negative; a long tail of zeros forces n positive.

✓ Worked example 1: the route on the map

$$0.00456 = 4.56 \times 10^{-3} \cdot 93,000,000 = 9.3 \times 10^7$$

given: two plain numbers · found: $n = -3$ and $n = +7$



🧠 **Sense check:** Same steps for both numbers. Only Step 3 differs: right for the small one, left for the large one. ✓

Scientific → standard: move the decimal back

Reverse it: the exponent says how many places to slide the decimal, and its sign says which way. Positive → right (bigger); negative → left (smaller).

$$9.3 \times 10^7 \rightarrow 93,000,000 \cdot 4.56 \times 10^{-3} \rightarrow 0.00456$$

+7: shift right 7 places, filling zeros · -3: shift left 3 places

Your turn: convert 0.00072

Fill each blank, then check.

step	value
coefficient M	_____
decimal moves	_____
exponent n	_____

$$0.00072 = 7.2 \times 10^{-4}$$

move right 4 places · number < 1 $\rightarrow n = -4 \cdot M = 7.2$



Worked example 2: entering 6.022×10^{23}

6.022×10^{23} on a calculator

given: 6.022×10^{23} · wanted: the keying that enters it as one number

Every mole calculation starts by keying this number in. The EE key (EXP on some models) means "times ten to the".

A common first attempt keys the $\times 10$ out loud: 6.022×10 EE 23. Test it.

✓ Worked example 2: solution

6.022×10^{23} on a calculator

EE means "times ten to the" · wanted: one number in the machine

A common first attempt

keyed 6.022×10 EE 23 → the machine holds $6.022 \times 10 \times 10^{23} = 6.022 \times 10^{24} \times$
EE already supplies the $\times 10$ · keying $\times 10$ again multiplies in an extra ten

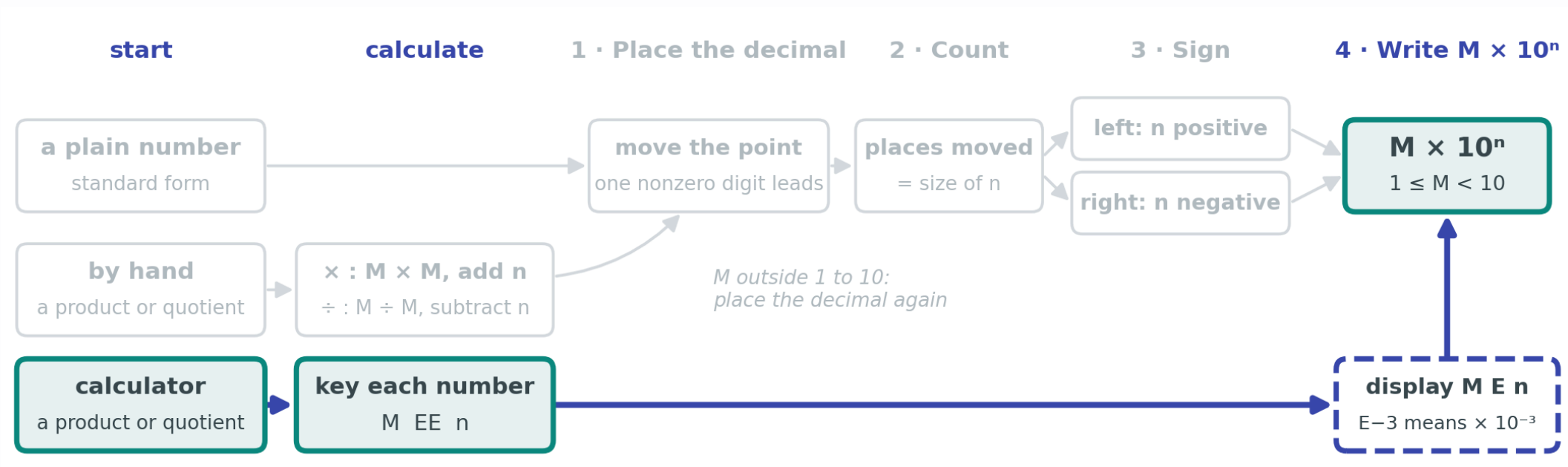
The correct keying

keyed 6.022 EE 23 → display $6.022\text{E}23 = 6.022 \times 10^{23}$ ✓
the whole number enters as one object: coefficient, then EE, then exponent

 **Sense check:** Read the display back before computing: E23 on the screen is $\times 10^{23}$. A result ten times too large marks the $\times 10$ -before-EE slip.

✓ Worked example 2: the route on the map

keyed 6.022 EE 23 → display 6.022E23 = 6.022×10^{23}
 given: 6.022×10^{23} · found: the keying 6.022 EE 23



🧠 **Sense check:** EE enters M and n together. Steps 1 to 3 never run: the display already reads $M \times 10^n$. ✓

⚠ Where this goes wrong

Sign backwards. A number below 1 gets a negative n: $0.00456 = 4.56 \times 10^{-3}$, not 10^3 .

The $\times 10$ -before-EE slip. On a calculator, 6.022 EE 23 is 6.022×10^{23} ; keying 6.022×10 EE 23 multiplies in an extra ten.

Off-by-one count. Count decimal moves, not zeros: 93,000,000 moves 7 places $\rightarrow 10^7$.

Practice 1

Which statement about writing 0.00042 in scientific notation is correct?

- A. The exponent is negative, because the number is less than 1 (the decimal moves right).
- B. The exponent is positive, because moving the decimal builds a larger coefficient.
- C. The exponent just equals the count of zeros, so every small number gives the same n.
- D. The coefficient is 0.42, leaving the decimal where it first lands.

✓ Practice 1 · answer: A

$$0.00042 = 4.2 \times 10^{-4}$$

less than 1 → decimal moves right 4 places → $n = -4$

A: a number below 1 always takes a negative exponent, and the decimal moves right to build 4.2. B flips the sign, C confuses zeros with places, D leaves M below 1.



Sense check: Direction sets the sign, distance sets the size, never the count of zeros.



Practice 2

one grain of fine sand: $2.0 \times 10^{-6} \text{ g}$

a pinch: 4.5×10^3 grains

A pinch of fine sand holds 4.5×10^3 grains, each of mass $2.0 \times 10^{-6} \text{ g}$. What is the mass of the pinch, in grams?

- A. 9.0×10^9
- B. 9.0×10^{-2}
- C. 9.0×10^{-3}
- D. 9.0×10^3

✓ Practice 2 · answer: C

keyed 2.0 EE –6 × 4.5 EE 3 → display 9.0E–3 = 9.0×10^{-3} g (answer C)

each number enters as one object · E–3 reads back as $\times 10^{-3}$

A keyed the –6 as +6: $2.0 \times 10^6 \times 4.5 \times 10^3 = 9.0 \times 10^9$. B keyed $\times 10$ before EE, an extra ten: $2.0 \times 10 \times 10^{-6} \times 4.5 \times 10^3 = 9.0 \times 10^{-2}$. D read the display back with the sign dropped: E–3 is 10^{-3} , not 10^3 .

 **Sense check:** A pinch of sand is a few milligrams: 9.0×10^{-3} g = 0.0090 g. ✓

Check yourself

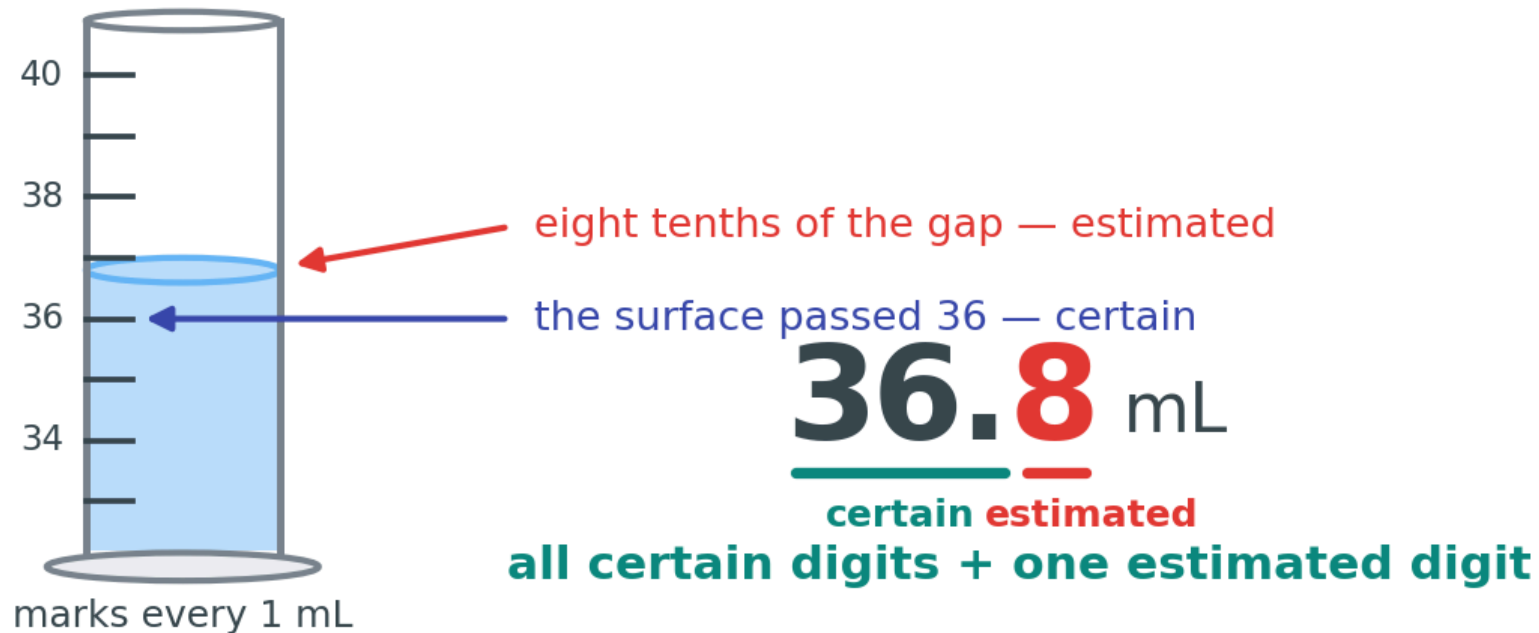
1. Write 0.000205 and 6,400,000 in scientific notation, and state the sign of each exponent.
2. Enter $(6.0 \times 10^{-3})(3.0 \times 10^5)$ on your calculator with the EE key, and write the display in proper scientific notation.

These powers of ten are the backbone of every mole, concentration, and atomic-mass calculation ahead.

7 • Significant Figures

Count the significant figures in any measurement, and round a calculated result with the rule that matches the operation.

Reading an instrument



The cylinder is marked every 1 mL. The water sits most of the way from 36 to 37: record 36.8 mL. The 3 and 6 are certain; the 8 is estimated.

A measurement's digits record its certainty

36.8 mL

3, 6 certain · 8 estimated: the true volume lies between 36.7 and 36.9 mL

An instrument reports every digit it can distinguish, plus one estimated digit. Those digits are the **significant figures**. Writing 36.8 mL states the volume is known to the tenths and no further.

A better instrument gives more digits

36.8 mL (graduated cylinder)
estimated in the tenths → 3 sig figs

36.82 mL (burette)
estimated in the hundredths → 4 sig figs

A burette marks every 0.1 mL, so its estimate lands in the hundredths. Recording 36.82 mL from a cylinder marked in whole milliliters claims a digit that was never measured.

Exact numbers

$$24 \text{ tablets} \cdot 1 \text{ kg} = 1000 \text{ g}$$

counted · defined → no estimated digit → exact, unlimited sig figs

24.31 g

measured: the final 1 is estimated → 4 sig figs

Counted objects and defined relationships are not measurements. Nothing in them is estimated, so they have unlimited significant figures. Only measured values limit a result.



The method

1. **Nonzero digits** always count.
2. **Captive zeros** (between nonzero digits) always count.
3. **Leading zeros** (before the first nonzero digit) never count.
4. **Trailing zeros** count only with a decimal point.



Worked example 1: a mass from the balance

0.04030 g

read from an analytical balance

An analytical balance reports the mass of a powder sample. Count the significant figures: test each digit against the four rules.

✓ Worked example 1: solution

0.04030 g

read from an analytical balance

Step 1 · Nonzero digits always count

Step 2 · Captive zeros always count

Step 3 · Leading zeros never count

Step 4 · Trailing zeros count only with a decimal point

0.04030 g → 4 sig figs

counted: 4, 0, 3, 0 · not counted: the two leading zeros

🧠 **Sense check:** In scientific notation the placeholders vanish: 4.030×10^{-2} g shows exactly the four significant digits.

✓ Worked example 1: the route on the map

0.04030 g

found: 4 sig figs · counted: 4, 0, 3, 0 · not counted: the two leading zeros

COUNT one number: test each digit

exact number counted · defined unlimited	nonzero digit 4.3 · 27 counts	captive zero 403 · 2.05 counts	leading zero 0.04 · 0.0072 never counts	trailing zero decimal point: 4.20 counts	trailing zero no decimal point: 420 placeholder
--	---	--	---	--	---

REPORT a calculated result



🧠 **Sense check:** All four digit rules ran on one number. Only the leading zeros were left out. ✓



Worked example 2: a scale with no decimal point

1200 kg

a truck-scale reading, no decimal point

A truck scale reports the mass of a loaded pallet. A common first attempt: four written digits, four significant figures. Test it against the rules.

✓ Worked example 2: solution

1200 kg

a truck-scale reading, no decimal point

Step 1 • Nonzero digits always count

Step 2 • Captive zeros always count

Step 3 • Leading zeros never count

Step 4 • Trailing zeros count only with a decimal point

1200 kg → 2 sig figs

counted: 1, 2 • placeholders: 0, 0 • the four-digit count fails Step 4

1200. kg → 4 sig figs • 1.20×10^3 kg → 3 sig figs

the decimal point makes trailing zeros count • the coefficient shows only significant digits

🧠 **Sense check:** Two sig figs: the thousands digit is certain, the hundreds digit is the estimate. Nothing in 1200 records the tens or ones.

✓ Worked example 2: the route on the map

1200 kg

found: 2 sig figs · counted: 1, 2 · placeholders: 0, 0

COUNT one number: test each digit

exact number counted · defined unlimited	nonzero digit 4.3 · 27 counts	captive zero 403 · 2.05 counts	leading zero 0.04 · 0.0072 never counts	trailing zero decimal point: 4.20 counts	trailing zero no decimal point: 420 placeholder
--	---	--	---	--	---

REPORT a calculated result



🧠 **Sense check:** With no decimal point, the trailing zeros take the placeholder branch. Only the nonzero digits count.



Take-home: trailing zeros need a decimal point

the same reading, three ways to write it

1200 kg

2 sig figs

no decimal point:
the zeros only hold place

1200. kg

4 sig figs

the decimal point makes
every trailing zero count

1.20 × **10³** kg

3 sig figs

the coefficient shows
only significant digits

■ significant ■ placeholder

1200 kg, 1200. kg, and 1.20×10^3 kg describe the same load with different certainty. To mark a trailing zero significant, write the decimal point or use scientific notation.

Your turn: count the sig figs

measurement	sig figs
0.0250 L	_____
305.0 g	_____
8700 m	_____

The decimal point decides every trailing zero.

0.0250 L → 3 · 305.0 g → 4 · 8700 m → 2

trailing zero after a decimal point counts · captive and trailing count · no decimal point: placeholders

⚠ Where this goes wrong

Counting trailing zeros with no decimal point. 2600 kg reads as four sig figs. Step 4 gives two: the zeros only hold place. Written 2600. kg, it has four.

Counting the leading zeros. In 0.0250 L, starting the count at the zeros gives 4 or 5. Leading zeros only locate the decimal point: 3 sig figs.

Dropping the trailing zero after a decimal point. The final zero of 0.0250 L was measured, and the decimal point makes it count: 3 sig figs, not 2.

Counting every written digit. 0.0250 L shows five digits but 3 sig figs. A digit is significant when it was measured, not when it is written.



Practice 1

0.02060 g

the mass of a grain of rice on an analytical balance

How many significant figures does the measurement carry?

- A. 3
- B. 4
- C. 5
- D. 6

✓ Practice 1 · answer: B

0.02060 g → 4 sig figs (answer B)

counted: 2, 0, 6, 0 · not counted: the two leading zeros

The 2 and 6 count, the captive zero between them counts, and the decimal point makes the final zero count. A dropped the trailing zero: 3. C started counting at the first zero after the decimal point: 5. D counted every written digit: 6.

 **Sense check:** Scientific notation strips the placeholders: 2.060×10^{-2} g keeps exactly four digits.

Rounding off

$$7.8342 \rightarrow 7.83$$

first dropped digit 4: below 5, the kept digit stays

$$0.4267 \rightarrow 0.43$$

first dropped digit 6: 5 or more, the kept digit rounds up

A calculator returns more digits than a measurement supports. Keep the significant ones and look at the first digit dropped: below 5, keep; 5 or more, round up.

Multiplication and division: fewest sig figs

$$4.20 \times 1.1 = 4.62 \rightarrow 4.6$$

3 sig figs $\times$ 2 sig figs $\rightarrow$ report 2 sig figs

A result can be no more certain than its least certain measurement. For multiplication and division, the answer keeps the **fewest sig figs** found among the inputs.



Worked example 3: volume of a block

$$8.5 \text{ cm} \times 4.27 \text{ cm} \times 1.36 \text{ cm}$$

given: three measured edges · wanted: the volume, correctly reported

A metal block's three edges are measured. Compute the volume and report it with the correct number of sig figs.

✓ Worked example 3: solution

$$8.5 \text{ cm} \times 4.27 \text{ cm} \times 1.36 \text{ cm}$$

sig figs: 2 · 3 · 3

Count each factor's sig figs

Multiply, then round to the fewest

$$8.5 \times 4.27 \times 1.36 = 49.3612 \text{ cm}^3 \text{ (calculator)} \rightarrow 49 \text{ cm}^3$$

reported to 2 sig figs, set by the 8.5

 **Sense check:** The 8.5 cm edge was estimated in the tenths. Two sig figs is everything those rulers measured; the calculator's extra digits were never measured at all.

✓ Worked example 3: the route on the map

$$8.5 \text{ cm} \times 4.27 \text{ cm} \times 1.36 \text{ cm} = 49 \text{ cm}^3$$

found: 49 cm^3 · 2 sig figs, set by the 8.5

COUNT one number: test each digit

exact number counted · defined unlimited	nonzero digit 4.3 · 27 counts	captive zero 403 · 2.05 counts	leading zero 0.04 · 0.0072 never counts	trailing zero decimal point: 4.20 counts	trailing zero no decimal point: 420 placeholder
--	---	--	---	--	---

REPORT a calculated result

count each factor



🧠 **Sense check:** Count each factor, keep the fewest sig figs, round once: $49.3612 \rightarrow 49 \text{ cm}^3$. The + or – rule stays unlit. ✓



Worked example 4: adding two volumes

$$108.2 \text{ mL} + 1.66 \text{ mL}$$

graduated cylinder (tenths) + burette (hundredths)

A burette adds 1.66 mL to the 108.2 mL already in a flask. A common first attempt: 1.66 has the fewest sig figs, three, so report 110. mL. Test it.

✓ Worked example 4: solution

$108.2 \text{ mL} + 1.66 \text{ mL}$
graduated cylinder (tenths) + burette (hundredths)

A common first attempt

$108.2 + 1.66 = 109.86 \rightarrow 110. \text{ mL}$
the multiplication rule applied to a sum ✗

Round to the least precise decimal place

$108.2 + 1.66 = 109.86 \rightarrow 109.9 \text{ mL}$
rounded to the tenths, the least precise place given ✓

 **Sense check:** Four sig figs survive even though one input carried three. Addition sets the answer's last decimal place; the sig-fig count follows from it.

✓ Worked example 4: the route on the map

$$108.2 \text{ mL} + 1.66 \text{ mL} = 109.9 \text{ mL}$$

found: 109.9 mL · rounded to the tenths, set by 108.2

COUNT one number: test each digit

exact number counted · defined unlimited	nonzero digit 4.3 · 27 counts	captive zero 403 · 2.05 counts	leading zero 0.04 · 0.0072 never counts	trailing zero decimal point: 4.20 counts	trailing zero no decimal point: 420 placeholder
--	---	--	---	--	---

REPORT a calculated result

count each factor

× or ÷ → **fewest sig figs**

+ or – → **least precise place**

both in one problem:
one rule per step

round once, at the end

dropped digit below 5: keep
dropped digit 5 or more: round up

🧠 **Sense check:** A sum takes the place rule, not the sig-fig count. The digit tests in the top row never decide it. ✓

Where the answer stops: the cutoff line

the cutoff

$$\begin{array}{r} 108.2 \\ + \quad 1.6 \\ \hline 109.8 \end{array}$$

108.2 stops at the tenths –
so the answer stops there too

→ 109.9 mL

rounded to the tenths: the last column left of the line

Stack the numbers at the decimal point. The least precise input draws a vertical line; nothing right of the line survives into the answer.

Match the rule to the operation

Do: for $\times$ and $\div$, keep the fewest sig figs.

$$8.5 \times 4.27 \times 1.36 = 49 \text{ cm}^3$$

fewest sig figs among the factors: two ✓

Do not: carry that rule into $+$ and $-$. Round to the least precise decimal place.

$$108.2 + 1.66 \rightarrow 110. \text{ mL } \times \cdot 109.9 \text{ mL } \checkmark$$

sig-fig count applied to a sum $\times \cdot$ a sum rounds to the tenths

Times counts digits; plus counts places.



Mixed operations: one rule per step

1. **Order of operations** decides the steps.
2. Each step's **own rule** decides what that intermediate is entitled to.
3. **Carry the digits unrounded**; round once, at the end, to the tightest entitlement.



Worked example 5: two rules in one calculation

$$(25.462 \text{ g} - 25.1 \text{ g}) \div 4.4 \text{ mL}$$

mass by difference from two balances · volume from a graduated cylinder

A vial is weighed full on an analytical balance, then empty on a coarser one; the sample it held is made up to 4.4 mL. Compute the mass per milliliter: one rule per step.

✓ Worked example 5: solution

$$(25.462 \text{ g} - 25.1 \text{ g}) \div 4.4 \text{ mL}$$

subtraction runs first: order of operations sets the steps

Subtract, and mark the entitlement

$$25.462 - 25.1 = 0.362 \text{ g} \rightarrow \text{entitled to the tenths (0.4-level)}$$

25.1 ends at the tenths $\rightarrow$ the difference holds one sig fig

Divide, then round once

$$0.362 \div 4.4 = 0.0823 \text{ g/mL (calculator)} \rightarrow 0.08 \text{ g/mL}$$

rounded once, at the end, to 1 sig fig, the tightest entitlement

 **Sense check:** Subtraction between near-equal readings destroys precision: the answer keeps one digit, though every input carried three or more.

✓ Worked example 5: the route on the map

$$(25.462 \text{ g} - 25.1 \text{ g}) \div 4.4 \text{ mL} = 0.08 \text{ g/mL}$$

found: 0.08 g/mL · 1 sig fig, set by the subtraction

COUNT one number: test each digit

exact number counted · defined unlimited	nonzero digit 4.3 · 27 counts	captive zero 403 · 2.05 counts	leading zero 0.04 · 0.0072 never counts	trailing zero decimal point: 4.20 counts	trailing zero no decimal point: 420 placeholder
--	---	--	---	--	---

REPORT a calculated result

count each factor

× or ÷ → **fewest sig figs**

+ or – → **least precise place**

both in one problem:
one rule per step

round once, at the end

dropped digit below 5: keep
dropped digit 5 or more: round up

 **Sense check:** Both rules ran, one per step: the place rule on the difference, the fewest-sig-figs rule on the division, one rounding at the end. ✓

Your turn: add three masses

$$14.55 \text{ g} + 0.322 \text{ g} + 2.1 \text{ g}$$

three balances, three precisions: where does the cutoff line fall?

Calculator: _____ → reported: _____

$$14.55 + 0.322 + 2.1 = 16.972 \rightarrow 17.0 \text{ g}$$

2.1 ends at the tenths → draw the cutoff line after the tenths



Sense check: Everything right of the line was never measured in the 2.1 g reading.

Your turn: multiply

$$3.10 \times 4.520$$

two measured values: count each factor's sig figs first

Calculator: _____ $\rightarrow$ reported: _____

$$3.10 \times 4.520 = 14.012 \rightarrow 14.0$$

fewest sig figs among the factors: three, from 3.10

 **Sense check:** The trailing zero in 14.0 counts: it is the third significant figure the 3.10 supports.

Your turn: subtract, then divide

$$(6.77 - 6.2) \div 2.33$$

subtraction first, division second: each step under its own rule

Difference: _____ $\rightarrow$ entitlement: _____ $\rightarrow$ reported: _____

$$6.77 - 6.2 = 0.57 \rightarrow \text{entitled to the tenths (0.6-level)}$$

one sig fig territory: carry 0.57 unrounded into the division

$$0.57 \div 2.33 = 0.2446 \rightarrow 0.2$$

rounded once, to 1 sig fig: the subtraction set the limit

 **Sense check:** The three sig figs in 2.33 never mattered; the subtraction's one-digit entitlement caps the answer.



Practice 2

salt: 20.07 g and 1.9 g, from two balances
all of it dissolved to make 150.0 mL of solution

Two portions of salt, **20.07 g** and **1.9 g**, are dissolved together to make **150.0 mL** of solution. What concentration of salt, in g/mL, should be reported?

- A. 0.147
- B. 0.15
- C. 0.146
- D. 22.0
- E. 0.146467

✓ Practice 2 · answer: C

$20.07 + 1.9 = 21.97 \text{ g} \rightarrow$ entitled to the tenths: 3 sig figs

sum: place rule, set by 1.9 · carry 21.97 unrounded

$21.97 \div 150.0 = 0.146467 \rightarrow 0.146 \text{ g/mL}$ (answer C)

division: fewest sig figs, 3 from the sum vs 4 from 150.0

A rounded the sum to 22.0 before dividing: $22.0 \div 150.0 = 0.1467 \rightarrow 0.147$. B used one rule throughout: 1.9 has 2 sig figs, so 0.15. D stopped at the sum: 22.0 g is the salt, not the salt per milliliter. E kept every calculator digit.

🧠 **Sense check:** The sum rule, not the sig-fig count of 1.9, sets the limit: 21.97 keeps three digits. ✓

Check yourself

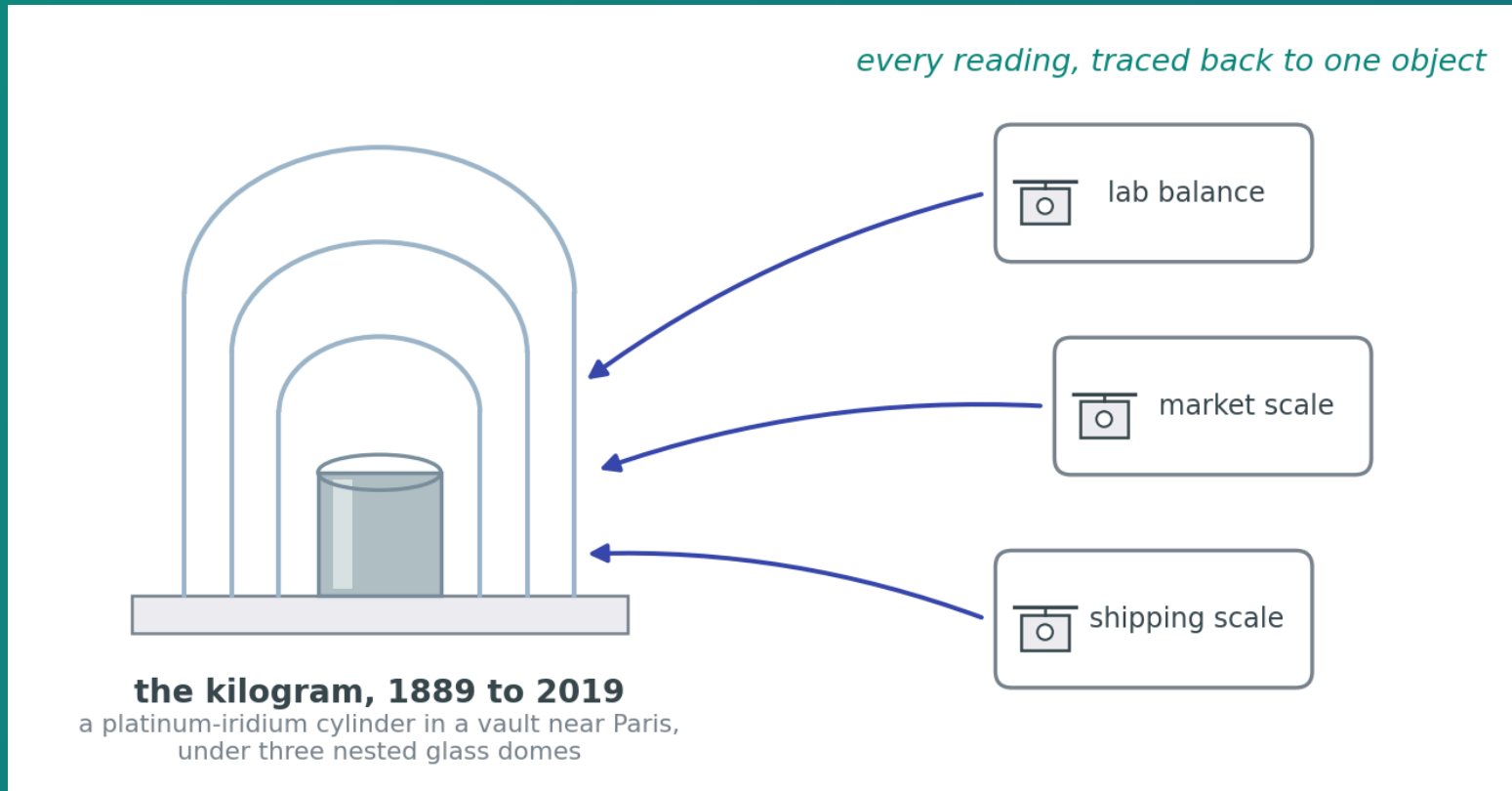
1. A balance reads **25.10 g**. How many sig figs, and which digit is the estimate?
2. 4.6×1.23 and $4.6 + 1.23$: which rule rounds each result, and to what?

These rules follow every measurement through every calculation. Unit conversions chain measurements with **exact** conversion factors: exact numbers never limit sig figs, so the measurement's certainty sets the answer's.

8 • The SI Unit System

Report every measurement as a number plus a unit, name the five SI base units chemistry uses, and read working and built units like the gram, the milliliter, and g/mL as combinations of them.

One object was the kilogram



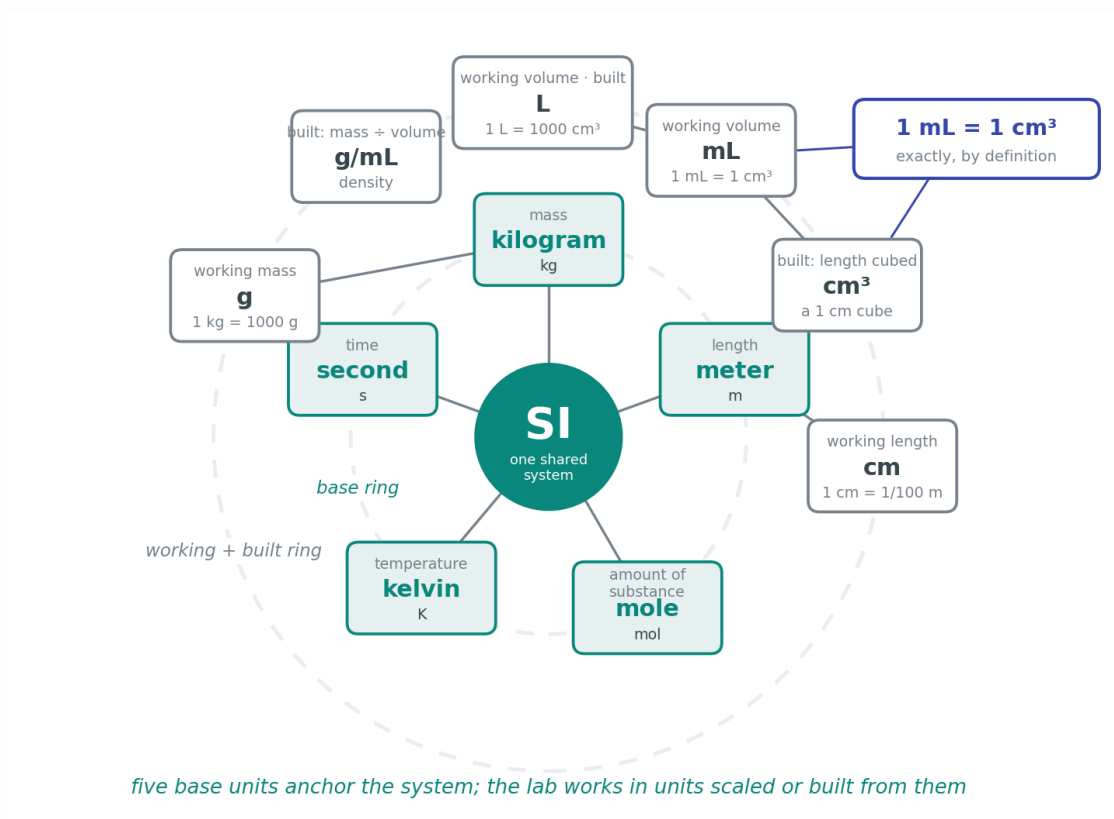
From 1889 to 2019, the kilogram was a platinum-iridium cylinder near Paris. Every balance on Earth traced to that one object. Three glass domes kept off the dust.

A measurement is a number and a unit

500 → not a measurement · 500 mg → one aspirin tablet
the unit names what was counted

A measured quantity reports two things. The number tells how much; the unit tells of what. Strip the unit and the number carries no information another scientist can use.

One system, five base units



The SI system gives every quantity one agreed base unit. Chemistry leans on five: the meter, the kilogram, the second, the kelvin, and the mole. Results measured anywhere compare directly.

Working units of the lab

$$1 \text{ kg} = 1000 \text{ g} \cdot 1 \text{ L} = 1000 \text{ mL}$$

balances read grams · glassware reads milliliters

A kilogram of reagent rarely sits on a bench. Balances read grams and glassware reads milliliters; each working unit ties to a base unit by a power of ten.

The liter is a built unit

$$10 \text{ cm} \times 10 \text{ cm} \times 10 \text{ cm} = 1000 \text{ cm}^3 = 1 \text{ L}$$

$$1 \text{ mL} = 1 \text{ cm}^3 \text{ exactly} \cdot 1 \text{ L} = 1000 \text{ mL}$$

Multiplying base units builds new ones. A cube 10 cm on each edge encloses one liter, and the milliliter and the cubic centimeter name the same volume exactly.

Density is a built unit

$$d = m / V \rightarrow \text{grams per milliliter} \rightarrow \text{g/mL}$$

two measurements combine into one new unit

Dividing a mass by a volume builds a new unit, the gram per milliliter. The unit itself states the meaning: the grams packed into each milliliter of a substance.



Practice 1

base units: $\text{m} \cdot \text{kg} \cdot \text{s} \cdot \text{K} \cdot \text{mol}$

wanted: the base unit for one measured quantity

A pharmacist weighs a powder sample. Which SI base unit matches the quantity being measured?

- A. the kilogram (kg)
- B. the gram (g)
- C. the liter (L)
- D. the meter (m)

✓ Practice 1 · answer: A

weighing → mass → kilogram (kg) (answer A)
the one base unit whose name carries a prefix

B is the working unit of the bench, not the base unit; 1 kg = 1000 g ties every gram reading to the kilogram. C measures volume, the space a sample fills. D measures length.

🧠 **Sense check:** Name the quantity first, then match the unit: a balance measures mass, and mass takes the kilogram.

✓



Practice 2

four notebook entries: a reading and a note on its unit
wanted: the entry that is right on every count

Which notebook entry is correct?

- A. one paper clip: 1 kg, recorded in the SI base unit of mass
- B. a warm water bath: 310 K, recorded in a built unit
- C. a flask of solution: 250 m, recorded as its volume
- D. a sugar solution: 1.20 g/mL, recorded in a built unit

✓ Practice 2 · answer: D

1.20 g/mL: a density, grams divided by milliliters (answer D)
the unit matches the quantity · the size fits a solution · built from two units

A has the right unit and an absurd size: 1 kg is a bag of sugar, and a paper clip is about a gram. B misnames the kelvin; it is one of the five base units. C reports a length; a volume takes mL or L.

 **Sense check:** Three tests per entry: the unit matches the quantity, the number fits the object, and base or built is named right. ✓



Check yourself

1. A notebook entry reads "volume of solution: 250." State what is missing, then supply a reasonable unit.
2. Classify each as base or built: the kelvin, the liter, the gram per milliliter.

Every metric prefix, tera through pico, defines an equality with a base unit. Chaining those equalities converts any measurement into any unit: that skill is dimensional analysis.

9 • Dimensional Analysis

Convert a measurement into any unit by chaining conversion factors, each one picked so the unit before it cancels.

Same number, wrong unit

same numbers, different units

the ground team reported

lbf·s · **pound-seconds**

≠

the navigation software read

N·s · **newton-seconds**

$$1 \text{ lbf} \cdot \text{s} = 4.448 \text{ N} \cdot \text{s}$$

the flyby that resulted

planned pass · 226 km

169 km
too low

actual pass · 57 km

atmosphere

Mars surface

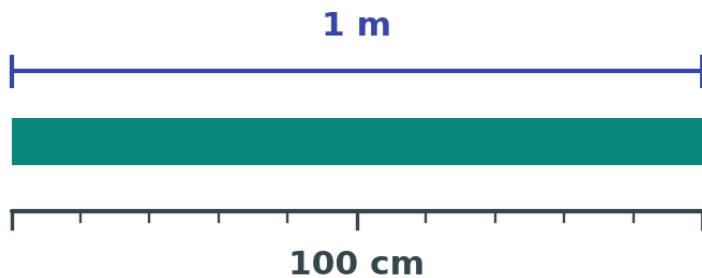
The Mars Climate Orbiter was lost in 1999. One team reported thruster impulse in **pound-seconds** and the software read **newton-seconds**. It flew too low and broke apart.

A conversion factor equals 1

$$1 \text{ m} = 100 \text{ cm}$$

one length, two names

An equality names one amount two ways. Written as a fraction, top matches bottom, so the fraction equals 1. Multiplying by 1 changes the unit, never the quantity.



one length, two names

$$\frac{1 \text{ m}}{100 \text{ cm}} = 1$$

$$\frac{100 \text{ cm}}{1 \text{ m}} = 1$$

top and bottom name
the same length

Units cancel like symbols in algebra

A unit on top cancels the same unit below. The right factor removes the given unit and leaves the wanted one.

$$\text{given } \cancel{\text{unit A}} \times \frac{\text{wanted unit B}}{\cancel{\text{unit A}}} = \text{answer, in unit B}$$

Metric prefixes are equalities

prefix	symbol	value
tera	T	10^{12}
giga	G	10^9
mega	M	10^6
kilo	k	10^3
hecto	h	10^2
deka	da	10^1

prefix	symbol	value
deci	d	10^{-1}
centi	c	10^{-2}
milli	m	10^{-3}
micro	μ	10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}

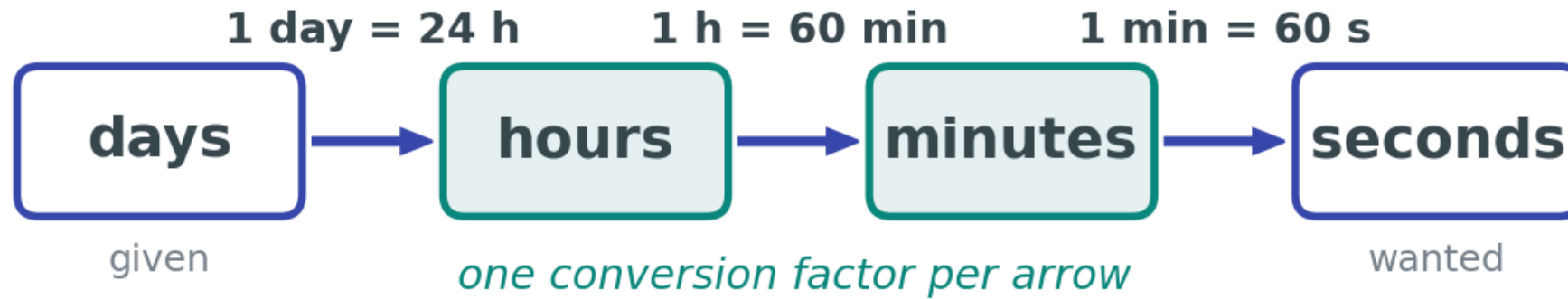
shaded rows appear in almost every problem

Each prefix defines an equality with any base unit: $1 \text{ km} = 10^3 \text{ m}$, $1 \text{ mg} = 10^{-3} \text{ g}$, $1 \text{ ms} = 10^{-3} \text{ s}$. One table covers every metric conversion.

King Henry Died By Drinking Chocolate Milk

kilo · hecto · deka · base · deci · centi · milli: one power of ten per word · beyond this run, the table steps by thousands

The unit plan



Sketch the route from the given unit to the wanted unit before any arithmetic. One conversion factor per arrow. When no single equality links them, the route runs through units in between.



The method

1. **Write the given:** number and unit.
2. **Pick the factor that cancels its unit:** the given unit goes in the denominator.
3. **Multiply; repeat until the wanted unit survives.**
4. **Sense-check** the size and the surviving unit.



Worked example 1

Step 1 • Write the given

$$1 \text{ m} = 100 \text{ cm}$$

given: 347 cm • wanted: m

A whiteboard measures **347 cm** across. Express the width in meters.

A common first attempt uses the factor written 100 cm over 1 m. Test it.

✓ Worked example 1: solution

1 m = 100 cm
given: 347 cm · wanted: m

A common first attempt

$$347 \text{ cm} \times \frac{100 \text{ cm}}{1 \text{ m}} = 34,700 \text{ cm}^2/\text{m} \times$$

Step 2 · Pick the factor that cancels its unit

Step 3 · Multiply; repeat until the wanted unit survives

$$347 \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}} = 3.47 \text{ m}$$

 **Sense check:** A meter holds 100 cm, so the count in meters must be smaller: 347 → 3.47. The width itself is unchanged. ✓

✓ Worked example 1: the route on the map

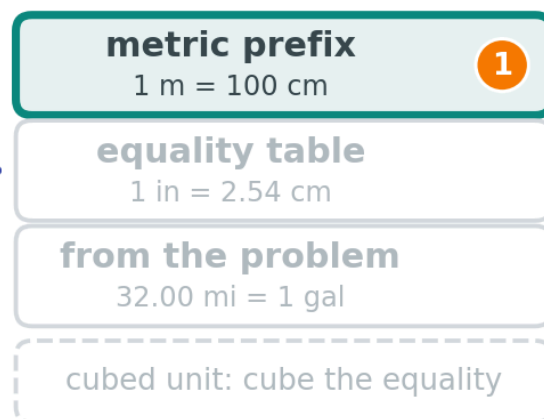
$$1 \text{ m} = 100 \text{ cm}$$

given: 347 cm · found: 3.47 m

1 · Write the given

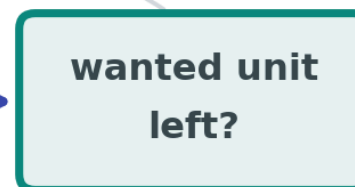


2 · Pick the factor

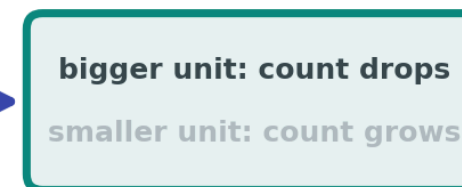


3 · Multiply; repeat

no: next factor



4 · Sense-check



 **Sense check:** One prefix factor, with cm on the bottom. m is the bigger unit, so the count dropped. ✓

Take-home: only one orientation cancels

Do: the given unit in the denominator, so it cancels.

$$347 \text{ cm} \times (1 \text{ m} / 100 \text{ cm}) = 3.47 \text{ m}$$

cm cancels · m survives ✓

Do not: the given unit on top. Nothing cancels; the factor is inverted.

$$347 \text{ cm} \times (100 \text{ cm} / 1 \text{ m}) = 34,700 \text{ cm}^2/\text{m}$$

no unit cancels → flip the factor ✗



Worked example 2

Step 1 • Write the given

$$1 \mu\text{m} = 10^{-6} \text{ m} \cdot 1 \text{ mm} = 10^{-3} \text{ m}$$

given: $45 \mu\text{m}$ • wanted: mm

A fine human hair measures **$45 \mu\text{m}$** across. Express the width in millimeters.

No single equality links μm to mm. The unit plan runs through the base unit: $\mu\text{m} \rightarrow \text{m} \rightarrow \text{mm}$.

✓ Worked example 2: solution

$$1 \mu\text{m} = 10^{-6} \text{ m} \cdot 1 \text{ mm} = 10^{-3} \text{ m}$$

given: 45 μm · wanted: mm · plan: $\mu\text{m} \rightarrow \text{m} \rightarrow \text{mm}$

Step 2 · Pick the factor that cancels its unit

$$45 \cancel{\mu\text{m}} \times \frac{10^{-6} \text{ m}}{1 \cancel{\mu\text{m}}} = 4.5 \times 10^{-5} \text{ m}$$

Step 3 · Multiply; repeat until the wanted unit survives

$$45 \cancel{\mu\text{m}} \times \frac{10^{-6} \cancel{\text{m}}}{1 \cancel{\mu\text{m}}} \times \frac{1 \text{ mm}}{10^{-3} \cancel{\text{m}}} = 0.045 \text{ mm}$$

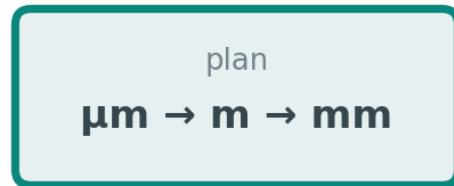
 **Sense check:** A millimeter holds 1000 μm , so the count drops by 1000: 45 $\rightarrow$ 0.045. A hair is thinner than a millimeter. ✓

✓ Worked example 2: the route on the map

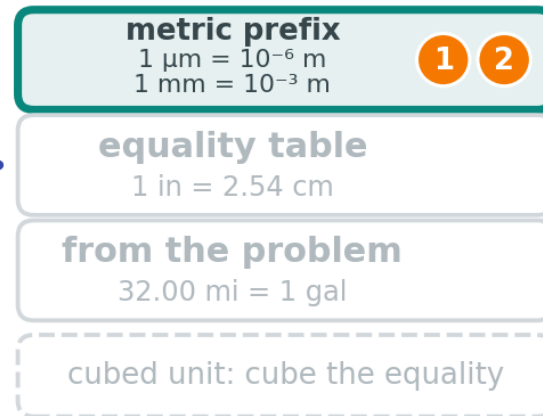
$$1 \mu\text{m} = 10^{-6} \text{ m} \cdot 1 \text{ mm} = 10^{-3} \text{ m}$$

given: $45 \mu\text{m}$ · found: 0.045 mm

1 · Write the given



2 · Pick the factor

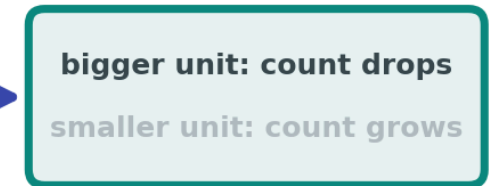


3 · Multiply; repeat

no: next factor



4 · Sense-check



🧠 **Sense check:** Two prefix factors, through the base unit. After the first, m was left, not the wanted mm, so the chain repeated once. ✓

Your turn: feet to centimeters

$$1 \text{ ft} = 12 \text{ in} \cdot 1 \text{ in} = 2.54 \text{ cm}$$

given: 6.00 ft · wanted: cm

A doorway stands **6.00 ft** tall. The unit plan: ft → in → cm.

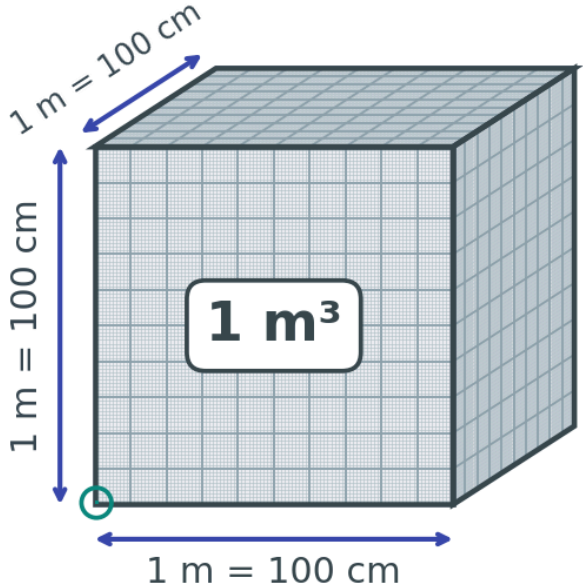
$$6.00 \text{ ft} \times \frac{12 \text{ in}}{1 \text{ ft}} \times \frac{\text{ } \text{cm}}{\text{ } \text{in}} = \text{ } \text{cm}$$

Fill the second factor from 1 in = 2.54 cm, then compute.

$$6.00 \text{ ft} \times \frac{12 \text{ in}}{1 \text{ ft}} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = \mathbf{183 \text{ cm}}$$

Cubed units: cube the whole equality

one cubic meter holds a million cubic centimeters



$(1\text{ m})^3 = (100\text{ cm})^3$

$1\text{ m}^3 = 100 \times 100 \times 100\text{ cm}^3$

$1\text{ m}^3 = 10^6\text{ cm}^3$

the number and the unit are both cubed

 one small cube = 1 cm^3

$100 \times 100 \times 100 = 1,000,000$ small cubes

Volume is length $\times$ length $\times$ length. A cubed unit takes the length equality cubed, number and unit together: $(1\text{ m})^3 = (100\text{ cm})^3$, so $1\text{ m}^3 = 10^6\text{ cm}^3$.



Worked example 3: cubic meters to cubic centimeters

Step 1 • Write the given

$$1 \text{ m} = 100 \text{ cm}$$

given: 3.5 m^3 • wanted: cm^3

A concrete pour measures 3.5 m^3 . Express the volume in cubic centimeters.

A common first attempt uses the linear factor, 100 cm over 1 m, once. Test it.

✓ Worked example 3: solution

$$1 \text{ m} = 100 \text{ cm}$$

given: 3.5 m^3 · wanted: cm^3

A common first attempt

$$3.5 \text{ m}^3 \times \frac{100 \text{ cm}}{1 \text{ m}} = 350 \text{ m}^2 \cdot \text{cm} \times$$

Step 2 · Pick the factor that cancels its unit

Step 3 · Multiply; repeat until the wanted unit survives

$$3.5 \text{ m}^3 \times \frac{10^6 \text{ cm}^3}{1 \text{ m}^3} = 3.5 \times 10^6 \text{ cm}^3$$

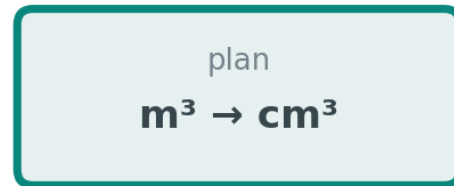
 **Sense check:** A centimeter cube is sugar-cube sized, and a million of them fill one cubic meter. The count grows by 10^6 : $3.5 \rightarrow 3,500,000$. The volume itself is unchanged. ✓

✓ Worked example 3: the route on the map

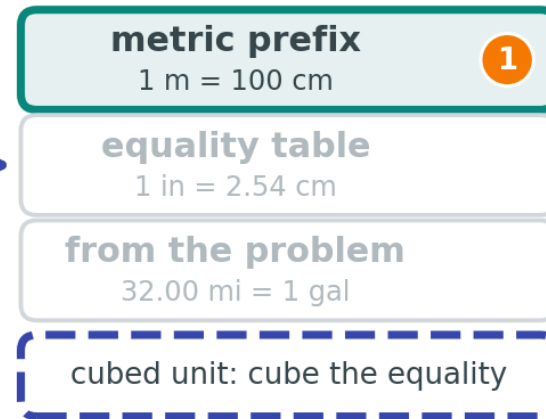
$$1 \text{ m} = 100 \text{ cm} \cdot (1 \text{ m})^3 = (100 \text{ cm})^3$$

given: 3.5 m^3 · found: $3.5 \times 10^6 \text{ cm}^3$

1 · Write the given

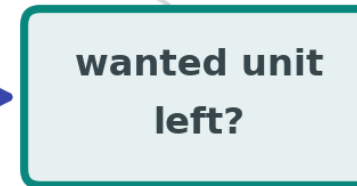


2 · Pick the factor

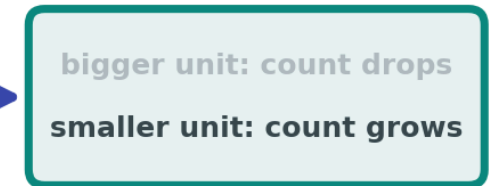


3 · Multiply; repeat

no: next factor



4 · Sense-check



🧠 **Sense check:** One prefix equality, cubed, number and unit together: $1 \text{ m}^3 = 10^6 \text{ cm}^3$. cm^3 is the smaller unit, so the count grew. ✓



Practice 1

$$1 \text{ m} = 100 \text{ cm}$$

given: $7.5 \times 10^5 \text{ cm}^3$ · wanted: m^3

An aquarium holds $7.5 \times 10^5 \text{ cm}^3$ of water. How many cubic meters does it hold?

- A. 7500
- B. 7.5×10^{11}
- C. 0.75
- D. 7.5×10^7

✓ Practice 1 · answer: C

$$1 \text{ m} = 100 \text{ cm} \cdot (1 \text{ m})^3 = (100 \text{ cm})^3$$

given: $7.5 \times 10^5 \text{ cm}^3$ · wanted: m^3

$$7.5 \times 10^5 \text{ cm}^3 \times \frac{1 \text{ m}^3}{10^6 \text{ cm}^3} = 0.75 \text{ m}^3 \text{ (answer C)}$$

A cubed the unit without cubing the number, dividing by only 100: $7.5 \times 10^5 / 100 = 7500$.
 B pointed the cubed factor the wrong way, so nothing cancels: $7.5 \times 10^5 \times 10^6 = 7.5 \times 10^{11}$.
 D used the linear factor, inverted: $7.5 \times 10^5 \times 100 = 7.5 \times 10^7$.

 **Sense check:** A cubic meter is a cube one meter on each edge, about refrigerator size. A large aquarium is somewhat smaller: 0.75 m^3 . The count drops by 10^6 from cm^3 to m^3 . ✓

Where this goes wrong

Inverting the factor. $347 \text{ cm} \times (100 \text{ cm} / 1 \text{ m}) = 34,700 \text{ cm}^2/\text{m}$. No unit cancels, and the answer is not in meters. If the units do not cancel, the factor is inverted: the correct setup gives 3.47 m.

Stopping mid-plan. The plan $\mu\text{m} \rightarrow \text{m} \rightarrow \text{mm}$ has two arrows. Stopping after one gives $4.5 \times 10^{-5} \text{ m}$: meters, not the wanted millimeters. The chain ends at 0.045 mm.

Calculating without a plan. On a multi-step chain, write the route first: given unit $\rightarrow \dots \rightarrow$ wanted unit. Each arrow names the factor to pick; a skipped arrow shows up as a unit that will not cancel.

Cubing the unit without cubing the number. $1 \text{ m}^3 = 100 \text{ cm}^3$ is false. Cubing $1 \text{ m} = 100 \text{ cm}$ cubes both sides entirely: $(1 \text{ m})^3 = (100 \text{ cm})^3 = 10^6 \text{ cm}^3$. The un-cubed factor turns 3.5 m^3 into 350 cm^3 instead of $3.5 \times 10^6 \text{ cm}^3$.

 Practice 2

$$1 \text{ mL} = 10^{-3} \text{ L} \cdot 1 \text{ gal} = 3.785 \text{ L}$$

given: 2500. mL · wanted: gal

A soft-drink bottle holds **2500. mL**. How many gallons does it hold?

- A. 2.500
- B. 0.6605
- C. 9.463
- D. 660.5

✓ Practice 2 · answer: B

$$1 \text{ mL} = 10^{-3} \text{ L} \cdot 1 \text{ gal} = 3.785 \text{ L}$$

given: 2500. mL · plan: mL → L → gal

$$2500. \cancel{\text{ mL}} \times \frac{10^{-3} \cancel{\text{ L}}}{1 \cancel{\text{ mL}}} \times \frac{1 \text{ gal}}{3.785 \cancel{\text{ L}}} = \mathbf{0.6605 \text{ gal (answer B)}}$$

Four sig figs survive: 2500. carries four, and the English-metric factor 3.785 counts too, so nothing limits below four. A stopped mid-plan: 2500. mL is 2.500 L, and liters are not gallons. C flipped the gallon factor: $2500./1000 \times 3.785 = 9.463$. D treated milliliters as liters: $2500./3.785 = 660.5$.

 **Sense check:** A gallon is nearly four liters. The bottle holds 2.500 L, less than one gallon: 0.6605. ✓



Worked example 4: a rate as a conversion factor

Step 1 • Write the given

$$1 \text{ mL} = 10^{-3} \text{ L} \cdot 1 \text{ gal} = 3.785 \text{ L} \cdot 0.62137 \text{ mi} = 1 \text{ km}$$

given: 1500. mL of gasoline • this car: 32.00 mi = 1 gal • wanted: km

A car gets **32.00 miles per gallon**. How many kilometers can it travel on **1500. mL** of gasoline?

Mileage is an equality for this car: 32.00 mi = 1 gal, a measured rating carrying four sig figs. The unit plan: mL → L → gal → mi → km.

✓ Worked example 4: solution

$$1 \text{ mL} = 10^{-3} \text{ L} \cdot 1 \text{ gal} = 3.785 \text{ L} \cdot 32.00 \text{ mi} = 1 \text{ gal} \cdot 0.62137 \text{ mi} = 1 \text{ km}$$

given: 1500. mL · plan: mL → L → gal → mi → km

Step 2 · Pick the factor that cancels its unit

$$1500. \cancel{\text{mL}} \times \frac{10^{-3} \text{ L}}{1 \cancel{\text{mL}}} = 1.500 \text{ L}$$

Step 3 · Multiply; repeat until the wanted unit survives

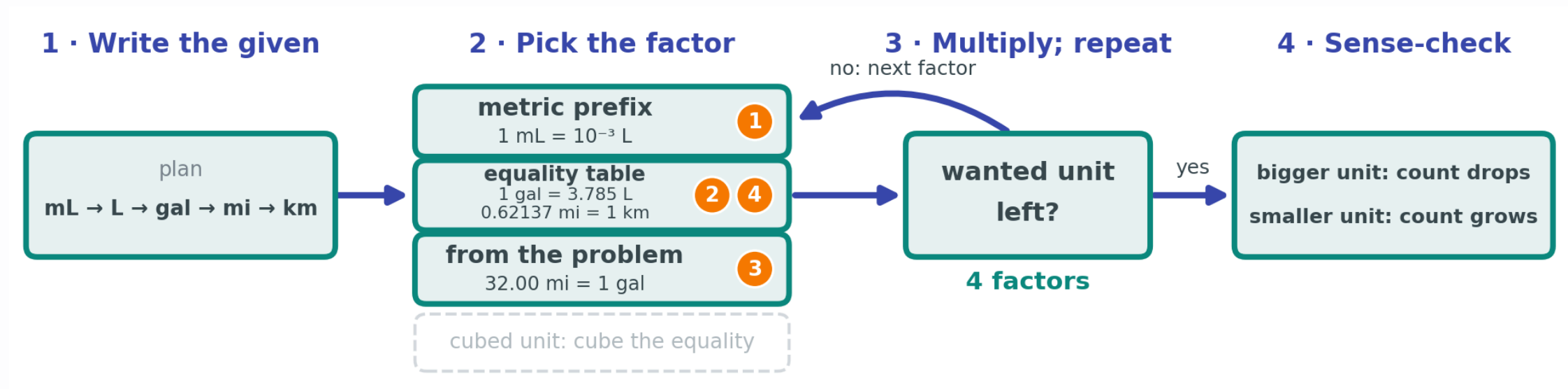
$$1500. \cancel{\text{mL}} \times \frac{10^{-3} \cancel{\text{L}}}{1 \cancel{\text{mL}}} \times \frac{1 \cancel{\text{gal}}}{3.785 \cancel{\text{L}}} \times \frac{32.00 \cancel{\text{mi}}}{1 \cancel{\text{gal}}} \times \frac{1 \text{ km}}{0.62137 \cancel{\text{mi}}} = 20.41 \text{ km}$$

 **Sense check:** 1500. mL is 1.5 L, under half a gallon: about 12.68 mi of driving. A kilometer is shorter than a mile, so the count in kilometers reads larger: 20.41. ✓

✓ Worked example 4: the route on the map

$$1 \text{ mL} = 10^{-3} \text{ L} \cdot 1 \text{ gal} = 3.785 \text{ L} \cdot 32.00 \text{ mi} = 1 \text{ gal} \cdot 0.62137 \text{ mi} = 1 \text{ km}$$

given: 1500. mL · found: 20.41 km



 **Sense check:** Four factors: a prefix, the gallon equality, the car's mileage, then the mile equality. Only the mileage comes from the problem; it holds for this car alone. ✓



Practice 3

$1 \text{ gal} = 3.785 \text{ L} \cdot \text{this hose: } 12.0 \text{ L} = 1 \text{ min} \cdot 1 \text{ min} = 60 \text{ s}$
given: 5.00 gal bucket · wanted: s to fill

A garden hose delivers **12.0 L per minute**. How many seconds does it take to fill a **5.00 gal** bucket?

- A. 6.60
- B. 25.0
- C. 1.58
- D. 94.6

✓ Practice 3 · answer: D

$$1 \text{ gal} = 3.785 \text{ L} \cdot 12.0 \text{ L} = 1 \text{ min} \cdot 1 \text{ min} = 60 \text{ s}$$

given: 5.00 gal · plan: gal → L → min → s

$$5.00 \text{ gal} \times \frac{3.785 \text{ L}}{1 \text{ gal}} \times \frac{1 \text{ min}}{12.0 \text{ L}} \times \frac{60 \text{ s}}{1 \text{ min}} = \mathbf{94.6 \text{ s (answer D)}}$$

The flow rate is a measured equality, 12.0 L = 1 min, and it converts liters to minutes. A flipped the gallon factor: 5.00 / 3.785 / 12.0 × 60 = 6.60. B skipped the gallon hop and treated gallons as liters: 5.00 / 12.0 × 60 = 25.0. C stopped at minutes: 5.00 × 3.785 / 12.0 = 1.58, a count of min, not s.

🧠 **Sense check:** 5.00 gal is about 19 L. At 12 L each minute that takes a bit over a minute and a half: 94.6 s. ✓

Check yourself

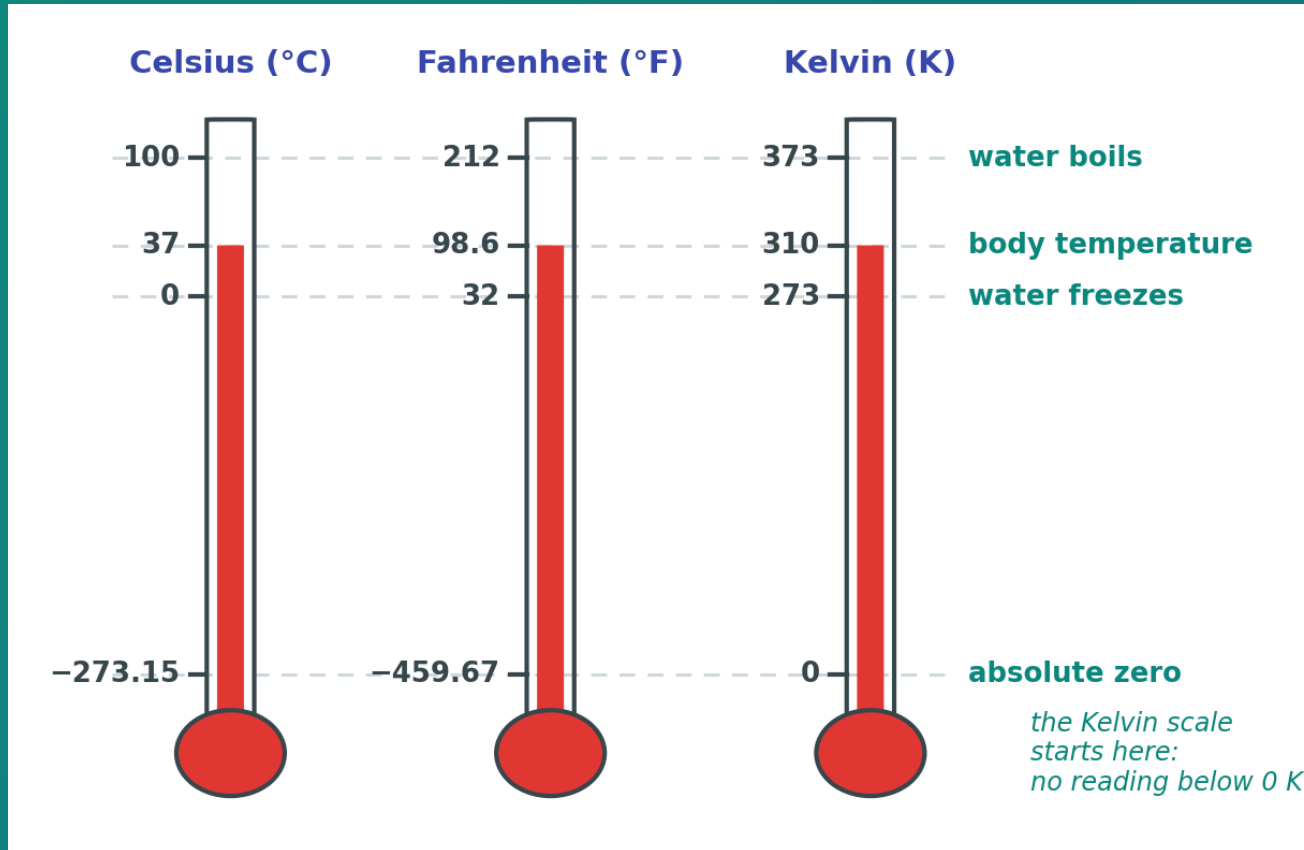
1. From **1 in = 2.54 cm**, write both conversion factors. Which one converts 30.0 cm to inches?
2. Multiplying by a conversion factor changes the unit but never the amount. What does every conversion factor equal?

Density is the next conversion factor: an equality between a substance's mass and its volume, in grams per milliliter. Molar mass and mole ratios follow. Every one converts on this same rail.

10 · Temperature Scales & Conversions

Convert a temperature reading among Celsius, Fahrenheit, and Kelvin, and handle a temperature difference correctly in any scale.

One temperature, three numbers



A Phoenix forecast says 95. A lab thermometer in the same air reads 35. A gas-law table lists the day as 308. Same afternoon, three scales.

A scale is two choices

water freezes: $0\text{ }^{\circ}\text{C} = 32\text{ }^{\circ}\text{F} = 273.15\text{ K}$

water boils: $100\text{ }^{\circ}\text{C} = 212\text{ }^{\circ}\text{F} = 373.15\text{ K}$

A temperature scale fixes two things: where its zero sits and how large one degree is. Celsius, Fahrenheit, and Kelvin mark the same physical events with different numbers.

Why Kelvin exists

$$0 \text{ K} = -273.15 \text{ }^{\circ}\text{C}$$

absolute zero: the coldest possible temperature · no negative kelvins

Kelvin puts its zero at absolute zero, where molecular motion reaches its minimum. Every Kelvin reading is positive. Gas volumes and pressures are proportional to Kelvin temperature, so the gas laws require this scale.

Celsius to Kelvin: shift the zero

$$K = ^\circ\text{C} + 273.15 \quad ^\circ\text{C} = K - 273.15$$

$$25\text{ }^\circ\text{C} + 273.15 = 298.15\text{ K} \quad 298.15\text{ K} - 273.15 = 25\text{ }^\circ\text{C}$$

A kelvin and a Celsius degree are the same size. Converting a reading is a single shift: add 273.15 going to Kelvin, subtract it coming back.

Celsius to Fahrenheit: stretch, then shift

$$^{\circ}\text{F} = 1.8(^{\circ}\text{C}) + 32 \quad \cdot \quad ^{\circ}\text{C} = (^{\circ}\text{F} - 32) / 1.8$$

$$\text{freezing to boiling: } 212 - 32 = 180 \text{ F}^{\circ} \text{ across } 100 \text{ C}^{\circ} \cdot 180/100 = 1.8$$

Fahrenheit degrees are smaller, so the reading is stretched by 1.8 before the zero shifts by 32. Going back reverses the order: subtract 32 first, then divide by 1.8.



The method

1. **Name the scales:** given and wanted.
2. **Pick the linking equation.**
3. **Solve, one operation at a time:** between $^{\circ}\text{F}$ and $^{\circ}\text{C}$, the 32 shifts last going in, first coming back.
4. **Sense-check** against the reference points.



Worked example 1

Step 1 • Name the scales

$$K = ^\circ C + 273.15$$

given: 37.0 °C • wanted: K

A water bath holds a protein sample at **37.0 °C**, body temperature. The instrument log records temperatures in kelvins. Express the reading in K.

✓ Worked example 1: solution

$$K = ^\circ\text{C} + 273.15$$

given: 37.0 °C · wanted: K

Step 2 · Pick the linking equation

Step 3 · Solve, one operation at a time

$$K = 37.0 + 273.15 = 310.15 \rightarrow 310.2 \text{ K}$$

one addition · reported to the tenths place of 37.0

Step 4 · Sense-check

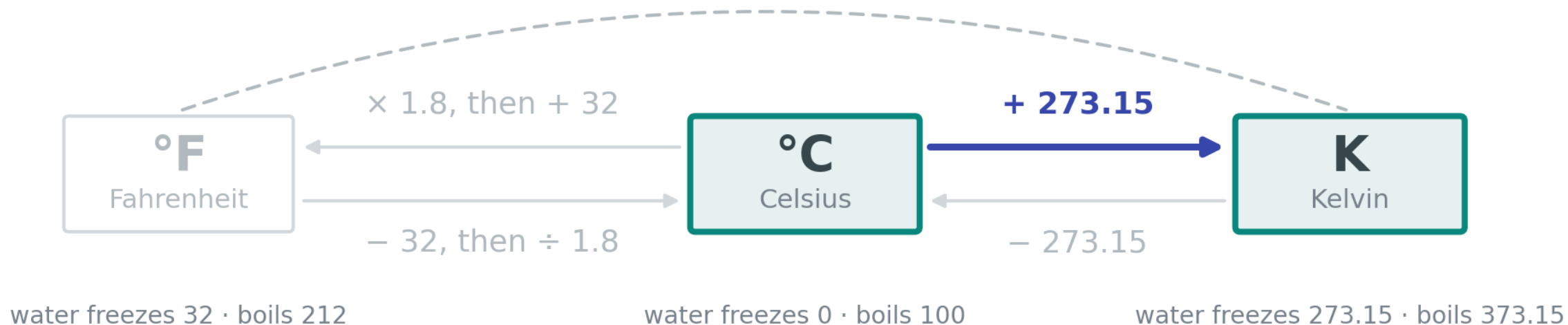
 **Sense check:** Body temperature sits between freezing (273.15 K) and boiling (373.15 K), nearer the freezing end: 310.2 K. ✓ The shift moved the number, not the warmth.

✓ Worked example 1: the route on the map

$$K = ^\circ C + 273.15$$

given: $37.0\ ^\circ C$ · found: $310.2\ K$

no direct arrow between $^\circ F$ and K : go through $^\circ C$



 **Sense check:** One arrow, Celsius to Kelvin: add 273.15. The degrees are the same size, so nothing stretches. ✓



Worked example 2

Step 1 • Name the scales

$$^{\circ}\text{F} = 1.8(^{\circ}\text{C}) + 32$$

given: 35.0°C • wanted: $^{\circ}\text{F}$

The Phoenix thermometer reads 35.0°C , and the forecast graphic wants Fahrenheit.

A common first attempt shifts before it stretches: add the 32 first. Test it.

✓ Worked example 2: solution

$$^{\circ}\text{F} = 1.8(^{\circ}\text{C}) + 32$$

given: 35.0 $^{\circ}\text{C}$ · wanted: $^{\circ}\text{F}$

A common first attempt

$$(35.0 + 32) \times 1.8 = 120.6 \text{ }^{\circ}\text{F} \times$$

the 32 got stretched along with the reading

Step 2 · Pick the linking equation

Step 3 · Solve, one operation at a time

$$^{\circ}\text{F} = 1.8(35.0) + 32 = 63.0 + 32 = 95.0 \text{ }^{\circ}\text{F}$$

$$\text{stretch: } 1.8 \times 35.0 = 63.0 \cdot \text{shift: } 63.0 + 32 = 95.0 \checkmark$$

Step 4 · Sense-check

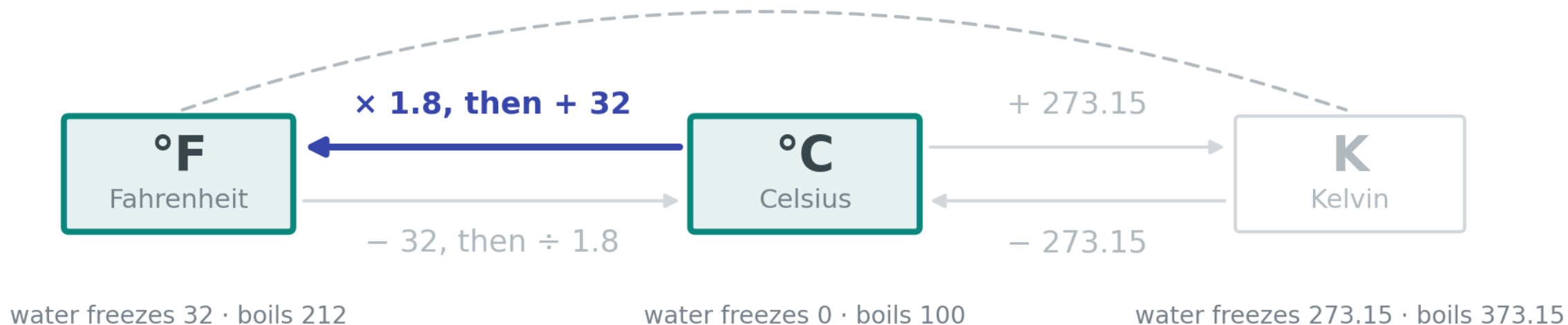
🧠 **Sense check:** 35.0 $^{\circ}\text{C}$ is just under body temperature (37 $^{\circ}\text{C}$), so $^{\circ}\text{F}$ must land just under 98.6: 95.0. ✓

✓ Worked example 2: the route on the map

$$^{\circ}\text{F} = 1.8(^{\circ}\text{C}) + 32$$

given: 35.0°C · found: 95.0°F

no direct arrow between $^{\circ}\text{F}$ and K : go through $^{\circ}\text{C}$



🧠 **Sense check:** One arrow, two operations in the arrow's order: $\times 1.8$ first, then $+ 32$. ✓

Take-home: the 32 never stretches

Do: multiply by 1.8 first, then add 32.

$$1.8(35.0) + 32 = 95.0 \text{ }^{\circ}\text{F}$$

stretch first · shift last ✓

Do not: add 32 first.

$$(35.0 + 32) \times 1.8 = 120.6 \text{ }^{\circ}\text{F}$$

the shift got stretched with the reading ✗

Your turn: Fahrenheit to Celsius

$$^{\circ}\text{C} = (^{\circ}\text{F} - 32) / 1.8$$

given: 68 $^{\circ}\text{F}$ · wanted: $^{\circ}\text{C}$

A thermostat is set to **68 $^{\circ}\text{F}$** . Coming back to Celsius, the order reverses: the 32 comes off first, then the 1.8 divides out.

$$^{\circ}\text{C} = (68 - 32) / 1.8 = \underline{\hspace{1cm}} / 1.8 = \underline{\hspace{1cm}} ^{\circ}\text{C}$$

Fill both blanks.

$$^{\circ}\text{C} = (68 - 32) / 1.8 = 36 / 1.8 = 20. ^{\circ}\text{C}$$

subtract first · divide last · a 20 $^{\circ}\text{C}$ room

A difference is not a reading

rise from 20 °C to 45 °C: $\Delta T = 45 - 20 = 25 \text{ C}^\circ = 25 \text{ K}$

the 273.15 shift cancels in the subtraction · in Fahrenheit: $1.8 \times 25 = 45 \text{ F}^\circ$

A reading marks a point; a difference measures a gap. Shifting both endpoints by 273.15 leaves the gap unchanged: a Celsius difference is a Kelvin difference, degree for degree.

Where this goes wrong

Adding 273.15 to a Fahrenheit reading. $98.6\text{ }^{\circ}\text{F} + 273.15 = 371.75$, a number on no scale. The 273.15 shift belongs to Celsius only. Route through Celsius: $(98.6 - 32) / 1.8 = 37.0\text{ }^{\circ}\text{C}$, then $37.0 + 273.15 = 310.15\text{ K}$.

Stretching by 1.8 on the way to Kelvin. $1.8(25) + 273.15 = 318.15$ is wrong; the 1.8 belongs to Fahrenheit. Celsius and Kelvin degrees are the same size: $25 + 273.15 = 298.15\text{ K}$.

Converting a temperature difference like a reading. A rise of $25\text{ }^{\circ}\text{C}$ is not a rise of 298.15 K . The shift cancels between the two endpoints: a $25\text{ }^{\circ}\text{C}$ rise is a 25 K rise, and $1.8 \times 25 = 45\text{ }^{\circ}\text{F}$.

 Practice 1

$$K = ^\circ\text{C} + 273.15$$

given: $-78\text{ }^\circ\text{C}$ · wanted: K

Dry ice sublimates at $-78\text{ }^\circ\text{C}$. What is this temperature in kelvins?

- A. 351
- B. -108
- C. -351
- D. 195

✓ Practice 1 · answer: D

$$K = -78 + 273.15 = 195.15 \rightarrow 195 \text{ K (answer D)}$$

–78 is known to the ones place · report 195 K

A dropped the minus sign: $78 + 273.15 = 351.15$. C subtracted the shift instead of adding it: $-78 - 273.15 = -351.15$, and negative kelvins do not exist. B built Fahrenheit instead of Kelvin: $1.8(-78) + 32 = -108.4$.

🧠 **Sense check:** Dry ice is cold, but absolute zero is far colder: 195 K sits 78 degrees below freezing (273.15 K) and well above 0 K. Any negative Kelvin answer fails before the arithmetic starts. ✓



Practice 2

$$^{\circ}\text{F} = 1.8(^{\circ}\text{C}) + 32 \cdot \text{K} = ^{\circ}\text{C} + 273.15$$

given: 223 K · wanted: $^{\circ}\text{F}$

A cold-test chamber holds a battery pack at **223 K**. Its door display reads in $^{\circ}\text{F}$. What should the display show?

- A. -58
- B. -50
- C. 4
- D. 122
- E. 433

✓ Practice 2 · answer: A

$^{\circ}\text{C} = 223 - 273.15 = -50.15 \rightarrow ^{\circ}\text{F} = 1.8(-50.15) + 32 = -58.27 \rightarrow -58 ^{\circ}\text{F}$ (answer A)
no direct K $\rightarrow$ $^{\circ}\text{F}$ equation: route through Celsius · shift, then stretch, then shift · ones place from 223

B stopped at Celsius: -50.15 , never stretched or shifted into $^{\circ}\text{F}$. C divided by 1.8, the coming-back move: $-50.15 / 1.8 + 32 = 4.14$. D dropped the minus sign: $1.8(50.15) + 32 = 122.27$. E fed kelvins straight into the Fahrenheit equation: $1.8(223) + 32 = 433.4$.

 **Sense check:** Below -40 the Fahrenheit number is the more negative one: $-50 ^{\circ}\text{C}$ lands at $-58 ^{\circ}\text{F}$. Far colder than a kitchen freezer, far above 0 K. ✓



Check yourself

1. Liquid nitrogen boils at 77 K. Convert the reading to $^{\circ}\text{C}$, and state why the answer must come out negative.
2. A reaction mixture warms from 22 $^{\circ}\text{C}$ to 47 $^{\circ}\text{C}$. Give the temperature change in $^{\circ}\text{C}$, in K, and in $^{\circ}\text{F}$.

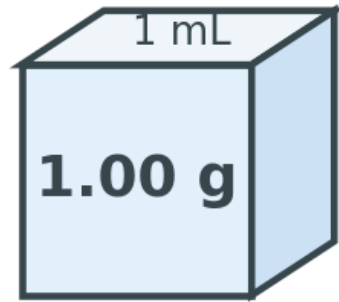
The gas laws ahead multiply and divide by temperature. Doubling a Kelvin temperature doubles a gas volume; doubling a Celsius reading means nothing. Every gas-law temperature enters in kelvins.

11 · Density as a Conversion Factor

Use a density to convert between the mass and the volume of a material, alone or chained with other conversion factors on one rail.

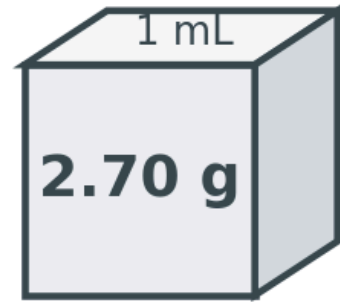
Same volume, different mass

the same milliliter, three different masses



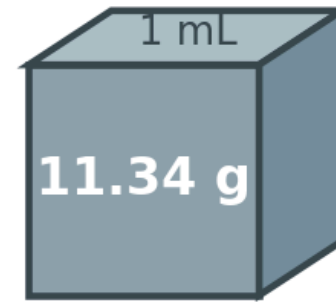
water

1.00 g/mL



aluminum

2.70 g/mL



lead

11.34 g/mL

Three cubes, the same 1 mL of space. Water: 1.00 g. Aluminum: 2.70 g. Lead: 11.34 g. Each material packs its own mass into a milliliter.

Density: mass per milliliter

ethanol 0.789 · water 1.00 · aluminum 2.70 · iron 7.87 · silver 10.5 · gold 19.3

density in g/mL: the mass, in grams, that fills 1 mL of the material

Each material packs a fixed mass into each milliliter. That rate, in grams per milliliter, is its **density**, a physical property. A measured density identifies the material.

A density is an equality

19.3 g of gold = 1 mL of gold
density of gold: 19.3 g/mL

Every equality gives a conversion factor. This one converts between mass and volume:

$$\frac{19.3 \text{ g Au}}{1 \text{ mL Au}} \text{ or } \frac{1 \text{ mL Au}}{19.3 \text{ g Au}}$$

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



The method

1. **Map the route:** given unit $\rightarrow$ desired unit.
2. **Start with the given.**
3. **Write the density fraction** so the unit to cancel sits in the denominator.
4. **Multiply and check** that only the desired unit survives.



Worked example 1: mass from volume

Step 1 • Map the route

$$13.6 \text{ g Hg} = 1 \text{ mL Hg}$$

given: 15.0 mL • wanted: g • route: mL → g

A sealed glass ampule holds **15.0 mL of mercury**. What is the mass of the mercury inside?

Set it up: which orientation of the density fraction cancels mL?

✓ Worked example 1: solution

$$13.6 \text{ g Hg} = 1 \text{ mL Hg}$$

given: 15.0 mL · wanted: g · route: mL → g

Step 2 · Start with the given

Step 3 · Write the density fraction

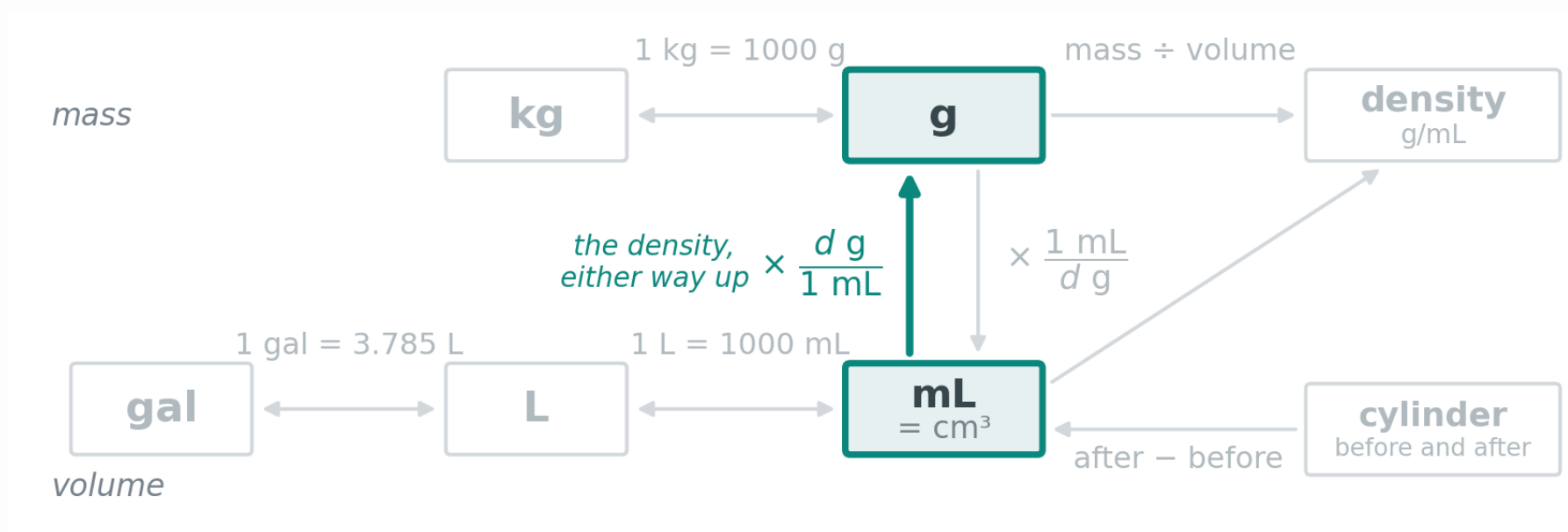
Step 4 · Multiply and check

$$15.0 \text{ mL Hg} \times \frac{13.6 \text{ g Hg}}{1 \text{ mL Hg}} = 204 \text{ g Hg}$$

 **Sense check:** Water in the same ampule would weigh 15.0 g. Mercury packs 13.6 times the mass into every milliliter: 204 g. ✓

✓ Worked example 1: the route on the map

13.6 g Hg = 1 mL Hg
given: 15.0 mL · found: 204 g



🧠 **Sense check:** One arrow, mL → g: the density with milliliters on the bottom. The given unit picks the orientation. ✓



Worked example 2: volume from mass

Step 1 • Map the route

$$19.3 \text{ g Au} = 1 \text{ mL Au}$$

given: 25.0 g · wanted: mL · route: g → mL

A pendant contains **25.0 g of gold**. What volume of gold is that?

A common first attempt: multiply by the density fraction as written, 19.3 g over 1 mL. Test the units.

✓ Worked example 2: solution

19.3 g Au = 1 mL Au
given: 25.0 g · wanted: mL · route: g → mL

A common first attempt

$$25.0 \text{ g Au} \times \frac{19.3 \text{ g Au}}{1 \text{ mL Au}} = 483 \text{ g}^2/\text{mL} \times$$

Step 2 · Start with the given

Step 3 · Write the density fraction

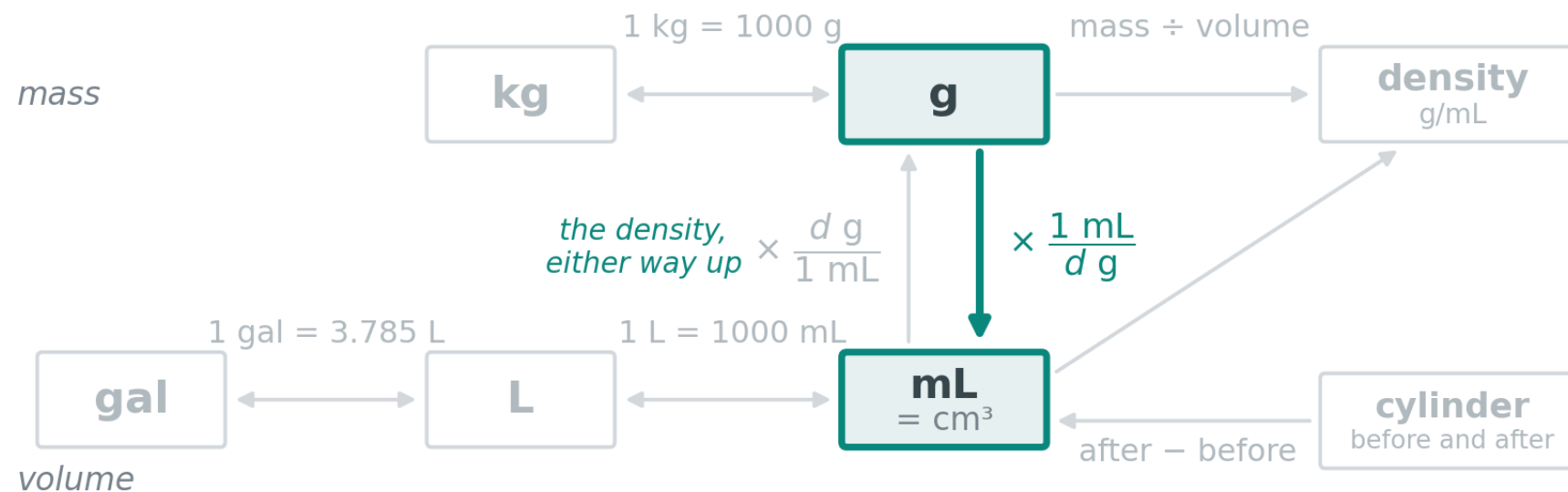
Step 4 · Multiply and check

$$25.0 \text{ g Au} \times \frac{1 \text{ mL Au}}{19.3 \text{ g Au}} = 1.30 \text{ mL Au}$$

🧠 **Sense check:** Gold packs 19.3 g into each milliliter, so 25.0 g fits in barely more than one: 1.30 mL. ✓

✓ Worked example 2: the route on the map

19.3 g Au = 1 mL Au
given: 25.0 g · found: 1.30 mL



🧠 **Sense check:** One arrow, g → mL: the density flipped, grams on the bottom. The given unit picks the orientation.

✓

Take-home: the units test the setup

$$25.0 \text{ g} \times \frac{19.3 \text{ g}}{1 \text{ mL}} = 483 \text{ g}^2/\text{mL} \times, \text{ not a volume}$$

$$25.0 \text{ g} \times \frac{1 \text{ mL}}{19.3 \text{ g}} = 1.30 \text{ mL} \checkmark$$

There is no multiply-or-divide rule to memorize. Write the fraction so the given unit cancels. A flipped fraction leaves units no quantity carries.

Your turn: ethanol

0.789 g ethanol = 1 mL ethanol

given: 50.0 g · wanted: mL

A hand-sanitizer recipe calls for **50.0 g of ethanol**, measured out by volume.

$$50.0 \text{ g} \times \frac{\text{_____ mL}}{\text{_____ g}} = \text{_____ mL}$$

Fill the fraction so grams cancel, then compute.

$$50.0 \text{ g} \times \frac{1 \text{ mL}}{0.789 \text{ g}} = \mathbf{63.4 \text{ mL}}$$

Where this goes wrong

Flipping the density fraction. $25.0 \text{ g} \times (19.3 \text{ g} / 1 \text{ mL}) = 483 \text{ g}^2/\text{mL}$. Nothing cancels, and no quantity carries g^2/mL . Put the unit to cancel in the denominator: $25.0 \text{ g} \times (1 \text{ mL} / 19.3 \text{ g}) = 1.30 \text{ mL}$.

Rearranging $d = m/V$ from memory. A misremembered rearrangement gives $13.6 \div 15.0 = 0.907 \text{ g/mL}^2$, which is not a mass. No rearranging is needed: start with the given and multiply by the fraction that cancels its unit.

A slipped decimal. A correct setup can still be keyed in wrong: 20.4 g instead of 204 g. Estimate first. 15 mL at about 14 g per milliliter is near 210 g, so 20.4 g cannot be right.



Practice 1

$10.5 \text{ g Ag} = 1 \text{ mL Ag}$
density of silver: 10.5 g/mL

A silversmith buys a **170. g silver ingot**. What volume, in mL, does the ingot occupy?

- A. 0.0618
- B. 16.2
- C. 1.62
- D. 1.79×10^3

✓ Practice 1 · answer: B

$$10.5 \text{ g Ag} = 1 \text{ mL Ag}$$

given: 170. g · wanted: mL · route: g → mL

$$170. \text{ g Ag} \times \frac{1 \text{ mL Ag}}{10.5 \text{ g Ag}} = 16.2 \text{ mL Ag (answer B)}$$

A divided the density by the mass: $10.5 \div 170. = 0.0618$, and its units are not milliliters. C slipped a decimal: $170. \div 10.5 = 16.2$, not 1.62. D flipped the fraction: $170. \times 10.5 = 1.79 \times 10^3$, with units of g^2/mL .

 **Sense check:** Silver packs 10.5 g into each milliliter, so 170. g occupies far fewer milliliters than its grams: 16.2. ✓



Worked example 3: an unknown metal

417 g of metal pellets · submerged, they raise the water level by 53.0 mL
given: 417 g and 53.0 mL · wanted: density, in g/mL

Candidate densities, in g/mL:

aluminum	iron	copper	silver	lead
2.70	7.87	8.96	10.5	11.34

A bin of unlabeled gray pellets arrives at a recycling yard. Density identifies the metal.

Build the density from the data, units in place, and match it to the table.

✓ Worked example 3: solution

417 g of metal pellets · 53.0 mL of water displaced
wanted: density, in g/mL

Build the fraction from the data

$$d = \frac{417 \text{ g}}{53.0 \text{ mL}} = 7.87 \text{ g/mL}$$

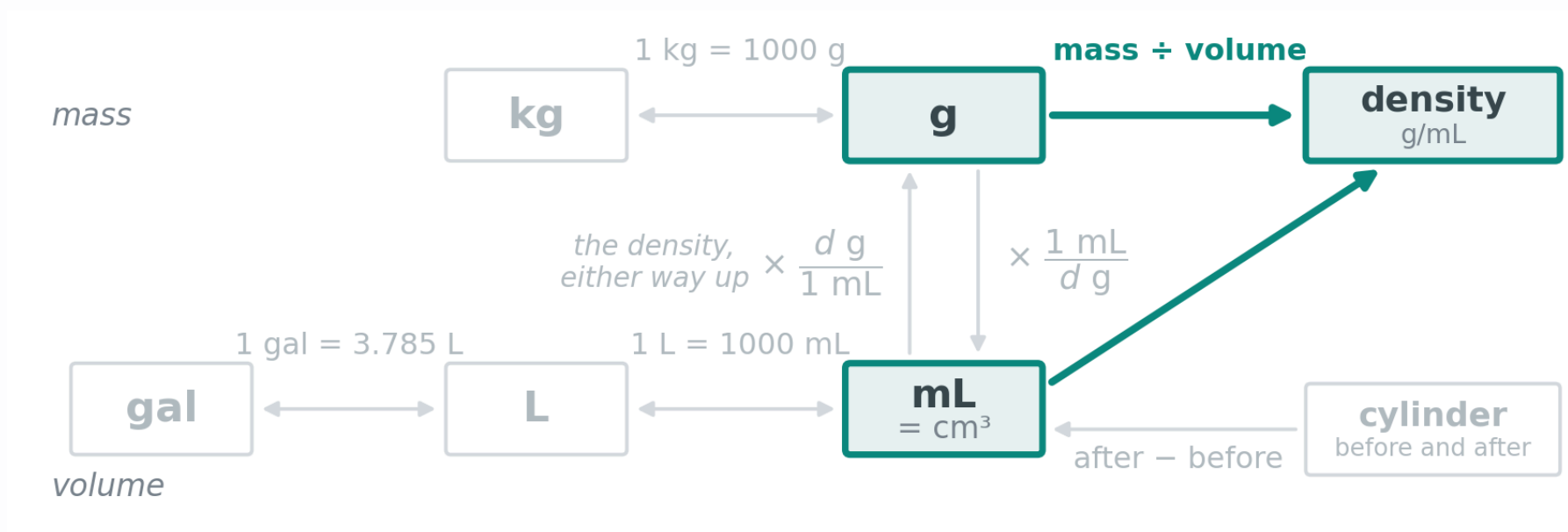
Match the property

aluminum	iron	copper	silver	lead
2.70	7.87	8.96	10.5	11.34

 **Sense check:** The measured fraction is now a conversion factor for these pellets: 7.87 g over 1 mL converts volume to mass; flipped, mass to volume. ✓

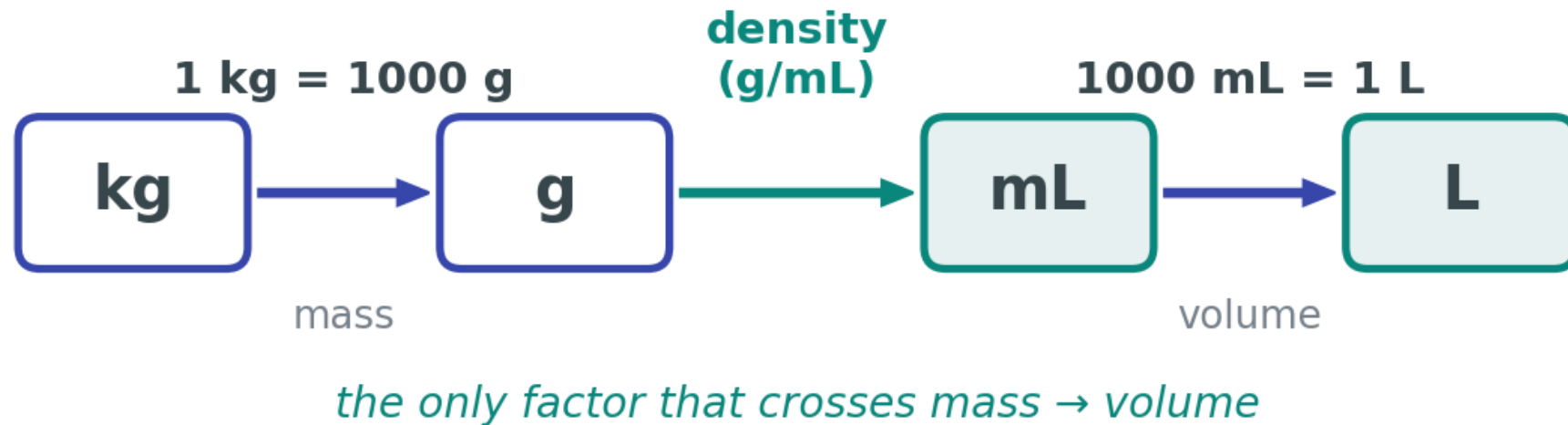
✓ Worked example 3: the route on the map

417 g of metal pellets · 53.0 mL of water displaced
found: 7.87 g/mL, iron



🧠 **Sense check:** No conversion this time. Mass ÷ volume builds the density from data in one move. ✓

Density inside a longer route



Metric factors convert within mass or within volume. The density is the only factor that crosses between them. Map the route, then chain the factors on one rail.



Worked example 4: kilograms to liters

Step 1 • Map the route

$$0.789 \text{ g ethanol} = 1 \text{ mL ethanol}$$

given: 2.50 kg • wanted: L • route: kg $\rightarrow$ g $\rightarrow$ mL $\rightarrow$ L

A stockroom order arrives: **2.50 kg of ethanol**. The flammables cabinet is labeled in liters. How many liters is this?

The density carries only the g $\rightarrow$ mL arrow. Metric equalities carry the other two.

✓ Worked example 4: solution

0.789 g ethanol = 1 mL ethanol

given: 2.50 kg · wanted: L · route: kg → g → mL → L

Step 2 · Start with the given

Step 3 · Write the density fraction

Step 4 · Multiply and check

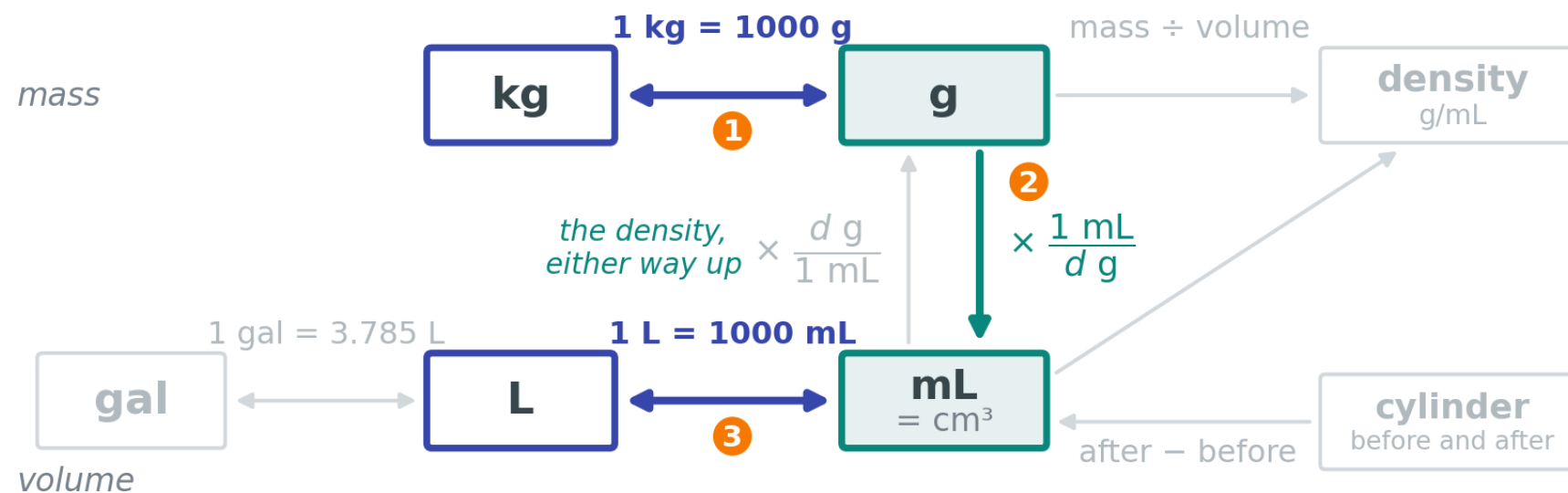
$$2.50 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mL}}{0.789 \text{ g}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 3.17 \text{ L}$$

 **Sense check:** Ethanol is lighter than water. Water would give exactly 2.50 L; ethanol spreads the same mass over 3.17 L. ✓

✓ Worked example 4: the route on the map

0.789 g ethanol = 1 mL ethanol

given: 2.50 kg · found: 3.17 L



🧠 **Sense check:** Three moves: kg → g, the density down to mL, then mL → L. Only the middle arrow crosses from mass to volume. ✓



Practice 2

$$1 \text{ gal} = 3.785 \text{ L} \cdot 1.84 \text{ g acid} = 1 \text{ mL acid}$$

given: 3.00 gal of concentrated sulfuric acid · wanted: kg

A stockroom receives **3.00 gal** of concentrated sulfuric acid, density **1.84 g/mL**. What mass, in kilograms, goes on the inventory sheet?

- A. 6.17
- B. 20,900
- C. 0.0209
- D. 20.9
- E. 1.46

✓ Practice 2 · answer: D

$$1 \text{ gal} = 3.785 \text{ L} \cdot 1.84 \text{ g acid} = 1 \text{ mL acid}$$

given: 3.00 gal · wanted: kg · route: gal → L → mL → g → kg

$$3.00 \text{ gal} \times \frac{3.785 \text{ L}}{1 \text{ gal}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1.84 \text{ g}}{1 \text{ mL}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 20.9 \text{ kg (answer D)}$$

A flipped the density: $11,355 \div 1.84 \div 1000 = 6.17$. B stopped at grams: 20,893 g is the mass in g, not kg. C skipped L → mL, applying g/mL to 11.355 L: $20.9 \text{ g} = 0.0209 \text{ kg}$. E flipped the gallon factor: $3.00 \div 3.785 \times 1000 \times 1.84 \div 1000 = 1.46$.

🧠 **Sense check:** Water would be 11.4 kg; the acid packs 1.84 times the mass per milliliter: 20.9 kg. ✓

Check yourself

1. Iron: 7.87 g/mL. State the equality this declares, then write the fraction that converts **grams of iron to milliliters**.
2. A route runs **kg → g → mL → L**. Which arrow is the density's, and why can no metric factor replace it?

Molar mass is the next factor of this kind: the grams in one **mole** of a substance. One fraction on the same rail converts a mass into a count of particles.

Can you...?

- ☐ define matter, place a description in the macroscopic, microscopic, or symbolic domain, and classify a statement as hypothesis, law, or theory?
- ☐ name the SI base units, tell a base unit from a derived unit, and pick a sensible metric unit for a measurement?
- ☐ classify matter as an element, a compound, or a mixture?
- ☐ distinguish precision from accuracy in measured data?
- ☐ report measurements and results with the correct significant figures?
- ☐ convert units with conversion factors, canceling units at each step?
- ☐ use density as a conversion factor between mass and volume?

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