

Reactions & Stoichiometry

CHEM 50 · Dr. Karmach

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

- Balance any chemical equation with the smallest whole-number coefficients
- Read a balanced equation as a recipe: coefficients are mole relationships
- Convert between amounts of any two species using mole ratios
- Chain molar masses with mole ratios to solve gram-to-gram problems
- Identify which reactant limits a reaction and how much product it allows
- Classify a reaction by type and cite the evidence that a reaction occurred
- Apply the solubility rules to predict precipitates and assign physical states
- Write molecular, complete ionic, and net ionic equations and classify electrolytes
- Assign oxidation numbers and identify oxidizing and reducing agents
- Write half-reactions and use the activity series to predict single displacement
- Carry molarity through stoichiometric calculations and compute percent yield
- Find an unknown concentration, mass percent, or formula from one measurement

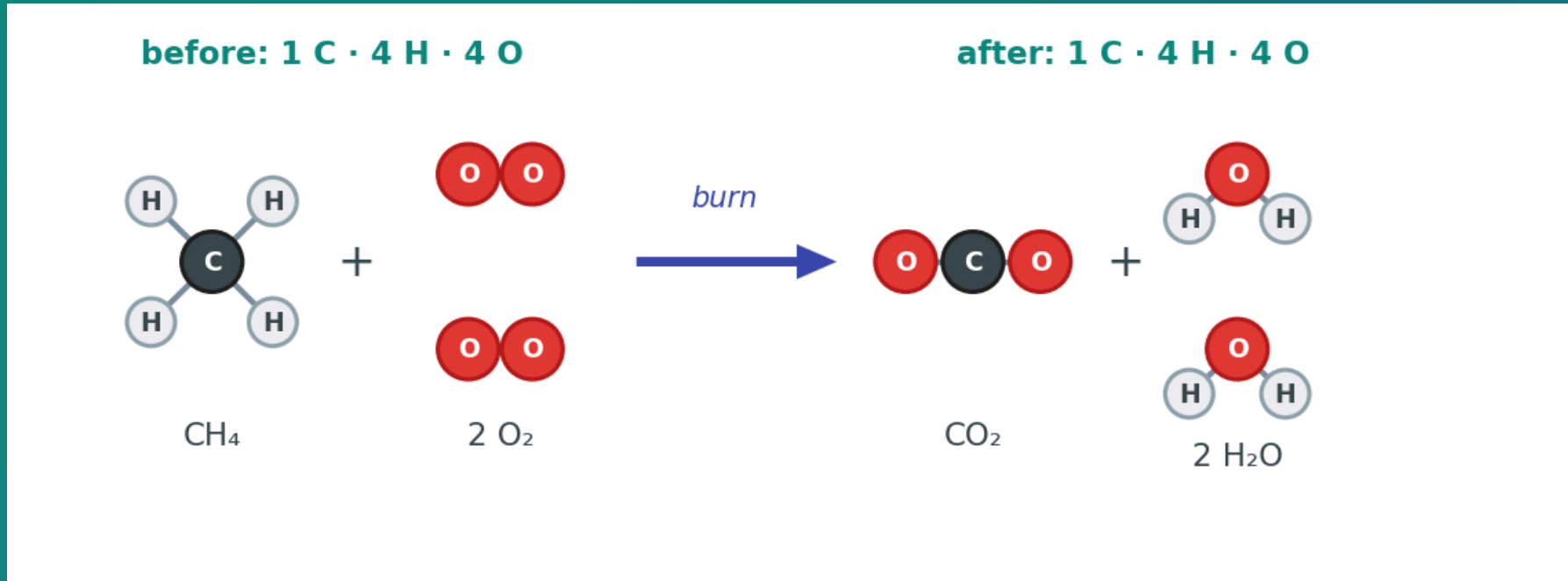
Today's route

1. **Balancing Equations**
2. **Types of Reactions**
3. **Solubility Rules & Precipitation**
4. **Net Ionic Equations**
5. **Electrolytes & Dissociation**
6. **Acids and Bases, Strong vs Weak**
7. **Oxidation Numbers**
8. **Oxidizing & Reducing Agents**
9. **The Activity Series**
10. **Mole Ratios**
11. **Mass-to-Mass Stoichiometry**
12. **Limiting Reactant**
13. **Percent Yield**
14. **Solution Stoichiometry**
15. **Quantitative Analysis**

1 • Balancing Equations

Turn any skeleton equation into a balanced one, and identify which numbers may change.

Atoms don't disappear



Burning methane makes new molecules out of the same atoms. Count them: 1 C, 4 H, 4 O before and after. Every reaction works this way.

What an equation says



reactants → products · the arrow reads "react to form" · each + reads "and"

"methane and oxygen react to form carbon dioxide and water"

(s) solid · (l) liquid · (g) gas · (aq) aqueous, dissolved in water

An equation is a sentence about substances. Formulas right, coefficients not yet chosen: that is a **skeleton equation**. Balancing supplies the coefficients; nothing else may change.

Your turn: count atoms in 3 Mg(OH)₂

Sixty seconds, straight from the naming rules: parentheses multiply, and now a coefficient multiplies the whole formula.

atom	count
Mg	$3 \times 1 = \underline{\hspace{1cm}}$
O	$3 \times \underline{\hspace{1cm}} = \underline{\hspace{1cm}}$
H	$3 \times \underline{\hspace{1cm}} = \underline{\hspace{1cm}}$



$$\text{Mg: } 3 \times 1 = 3 \cdot \text{O: } 3 \times 2 = 6 \cdot \text{H: } 3 \times 2 = 6$$

The coefficient 3 multiplies every atom in the formula; the subscript 2 multiplies only what the parentheses enclose.

Why every equation must balance

A reaction rearranges atoms. It never creates or destroys them. Both sides of a correct equation must show the **same number of each kind of atom**.



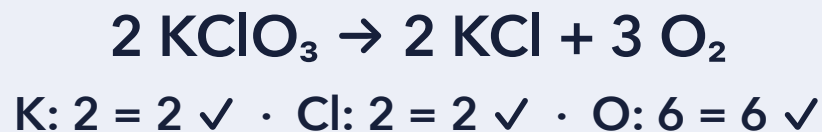
H: $4 = 4$ ✓ · O: $2 = 2$ ✓ · a possible reaction



H: $2 = 2$ ✓ · O: $2 \neq 1$ ✗ · an O atom would have to vanish

What "balanced" looks like

every atom count matches: the beam stays level



Every element's count matches. That is the complete test.



The method

1. Write the atom count for each side.
2. Balance **one element at a time**. Save any element that appears in several formulas for last.
3. **Recheck every count** after each change.
4. Finish with the smallest whole numbers.



Worked example 1: splitting water

Step 1 • Write the atom count for each side



H: 2 = 2 ✓ · O: 1 ≠ 2 ✗

Electrolysis splits water into hydrogen gas and oxygen gas. Oxygen doesn't balance.

A common first attempt: change H_2O to H_2O_2 . Test it.

✓ Worked example 1: solution



H: $2 = 2$ ✓ · O: $1 \neq 2$ ✗

A common first attempt



H: $2 = 2$ ✓ · O: $2 = 2$ ✓

Step 2 · Balance one element at a time



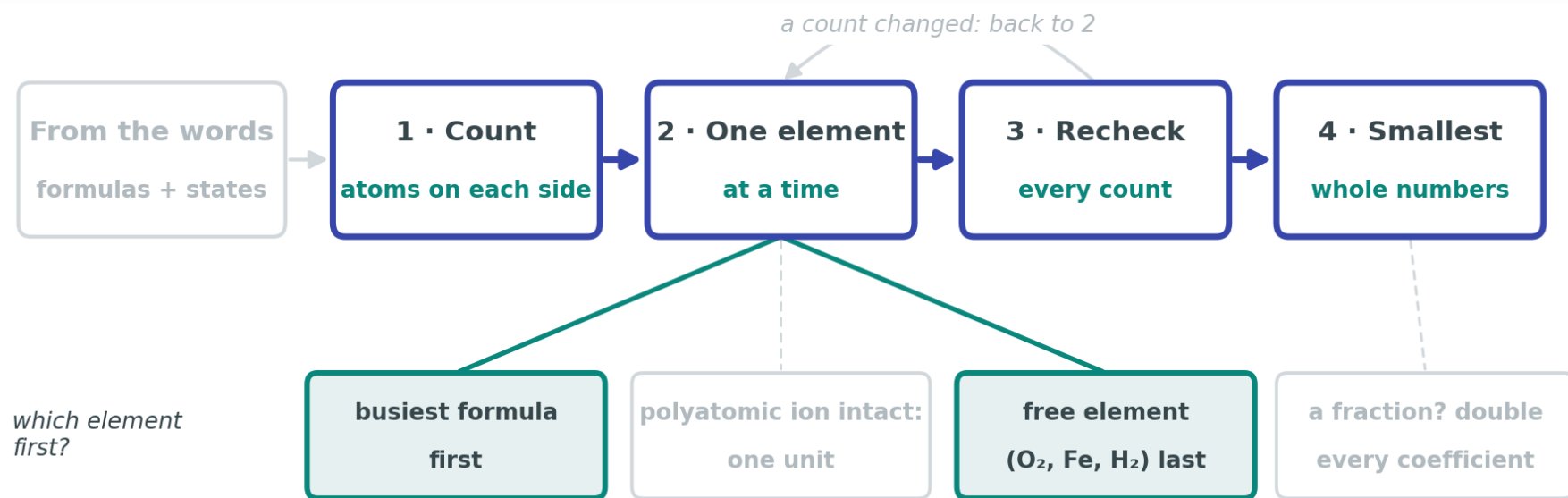
H: $4 = 4$ ✓ · O: $2 = 2$ ✓

 **Sense check:** Coefficients change the amount. Subscripts change the substance. Only coefficients may change.

✓ Worked example 1: the route



found: H: 4 = 4 ✓ · O: 2 = 2 ✓



 **Sense check:** H_2O is the only compound, so O went first: $2 \text{H}_2\text{O}$. H_2 , a free element, took the last coefficient: 2H_2 . The recheck found every count equal, so no step repeated.



Worked example 2: a common multiple

Rust: iron reacting with oxygen.

Step 1 • Write the atom count for each side



Fe: 1 $\neq$ 2 · O: 2 $\neq$ 3

Oxygen shows 2 on the left and 3 on the right. What is the smallest number both divide into?

✓ Worked example 2: solution



Fe: 1 $\neq$ 2 · O: 2 $\neq$ 3

Step 2 · Balance one element at a time



O: 6 = 6 ✓ · Fe: 1 $\neq$ 4 ✗

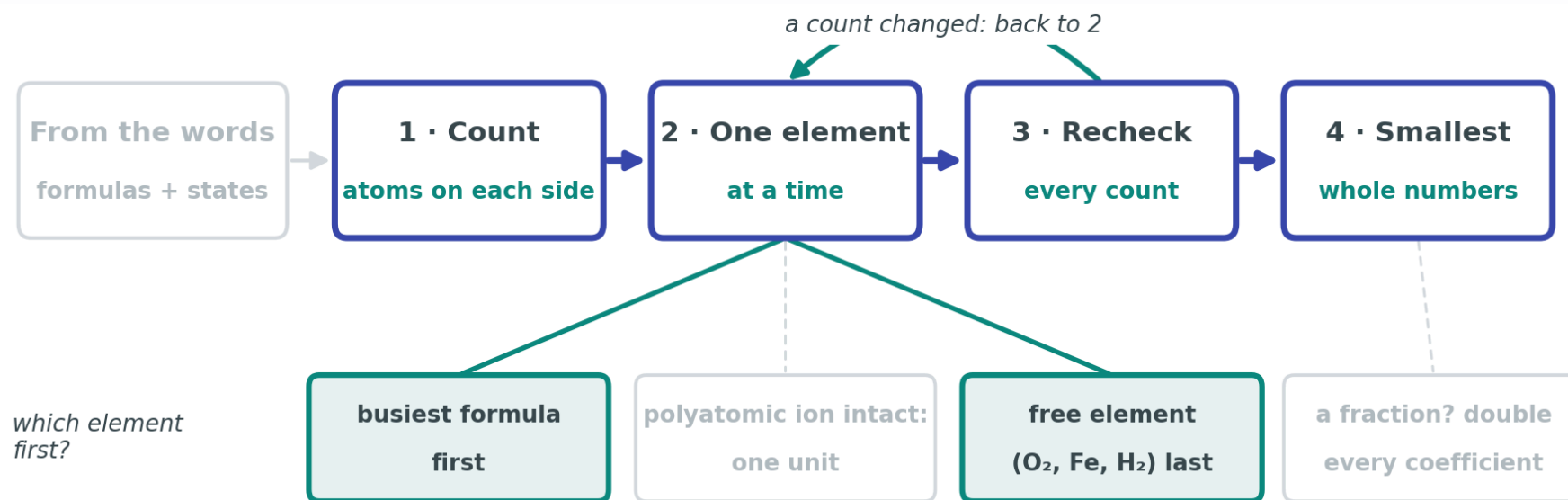
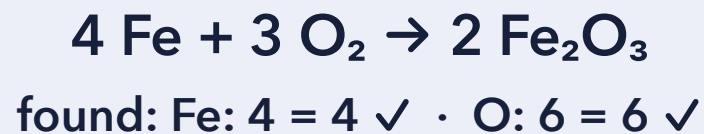
Step 3 · Recheck every count



Fe: 4 = 4 ✓ · O: 6 = 6 ✓

 **Sense check:** One change can unbalance an element already counted. Recheck every count after each change.

✓ Worked example 2: the route



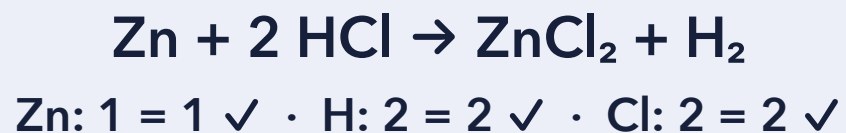
 **Sense check:** Fe₂O₃ is the busiest formula, so O went first, with the common multiple 6. The recheck caught Fe, and the free element took the last coefficient: 4 Fe.

Your turn: zinc and hydrochloric acid



atom	left	right
Zn	1	1 ✓
H	<u> </u> × 1	2
Cl	<u> </u> × 1	2

One coefficient balances both H and Cl.



Where this goes wrong

Changing subscripts. Writing H_2O_2 balances the counts but changes the substance. If the formula changes, the chemistry changes.

Forgetting the unwritten 1. $\text{Zn} + 2 \text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$ has coefficient sum $1 + 2 + 1 + 1 = 5$, not 4. An unwritten coefficient is still a 1.

Reading coefficients as grams. $4 \text{Fe} + 3 \text{O}_2$ does not mean "4 g and 3 g." Coefficients count particles or moles. Converting to mass requires the molar mass.

Stopping too early. $8 \text{Fe} + 6 \text{O}_2 \rightarrow 4 \text{Fe}_2\text{O}_3$ balances, but every coefficient divides by 2. Reduce to smallest whole numbers.



Practice 1



Ammonia burned over a platinum catalyst, the first step in making nitric acid for fertilizer. What does the **5** in front of O_2 mean?

- A. Each O_2 molecule contains 5 oxygen atoms.
- B. 5 molecules of O_2 react, or equally 5 moles, in proportion to the other coefficients.
- C. 5 grams of O_2 are consumed.
- D. Exactly 5 individual molecules react. The number cannot scale up to moles.

✓ Practice 1: answer B



$$5 \text{ O}_2 \rightarrow 5 \times 2 = 10 \text{ O} \cdot \text{ right side: } 4 \text{ NO} + 6 \text{ H}_2\text{O} \rightarrow 4 + 6 = 10 \text{ O} \checkmark$$

A describes a subscript's job. C: coefficients are counts, never masses. D: multiplying every coefficient by Avogadro's number leaves the ratios unchanged, so coefficients count moles as well as molecules.

🧠 **Sense check:** In 5 O_2 , the coefficient and the subscript do different jobs: $5 \times 2 = 10$ atoms of O.

Worked example 3: combustion

Stove gas is methane, and burning it makes carbon dioxide and water. The skeleton:

Step 1 • Write the atom count for each side



C: 1 = 1 ✓ · H: 4 ≠ 2 · O: 2 ≠ 3

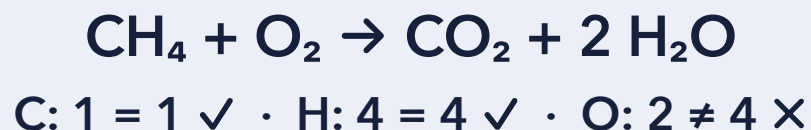
For combustion, balance **C first, H second, O last**. O₂ contains only one element, so its coefficient can be set last without disturbing the others.

Balance it.

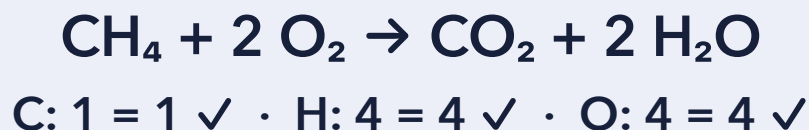
✓ Worked example 3: solution



Step 2 • Balance one element at a time

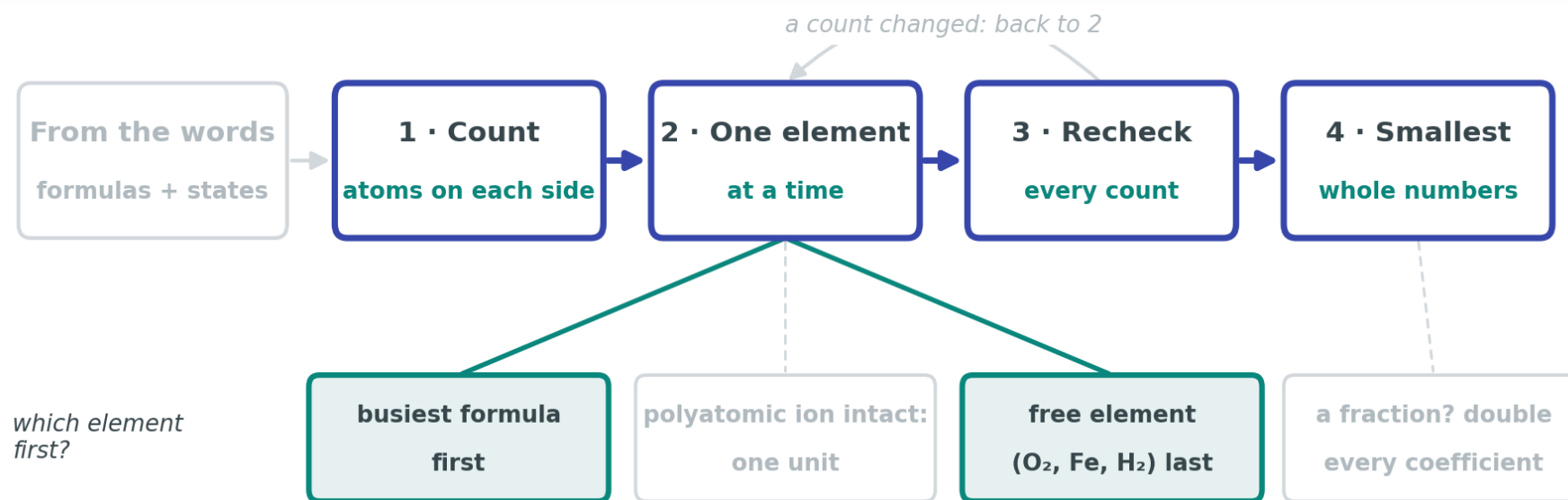
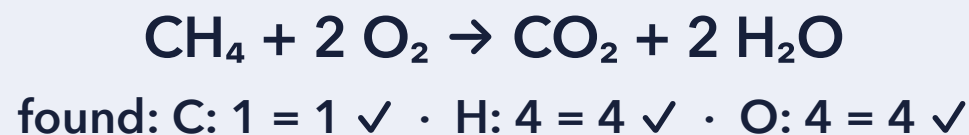


Step 3 • Recheck every count



 **Sense check:** 4 is even, so O₂ lands cleanly: no fractions, smallest whole numbers on the first pass. Saving the lone-element formula for last meant one clean choice at the end.

✓ Worked example 3: the route



 **Sense check:** CH₄ is the busiest formula: C, then H. O₂ is a free element, so its coefficient came last and changed no other count. Nothing sent the solution back.



Worked example 4: a fraction appears

Butane fuels lighters.

Step 1 • Write the atom count for each side



C: 4 ≠ 1 · H: 10 ≠ 2 · O: 2 ≠ 3

Same order: **C first, H second, O last.**

Balance it. A fraction will appear along the way.

✓ Worked example 4: solution



Step 2 • Balance one element at a time

Step 3 • Recheck every count

Step 4 • Finish with the smallest whole numbers

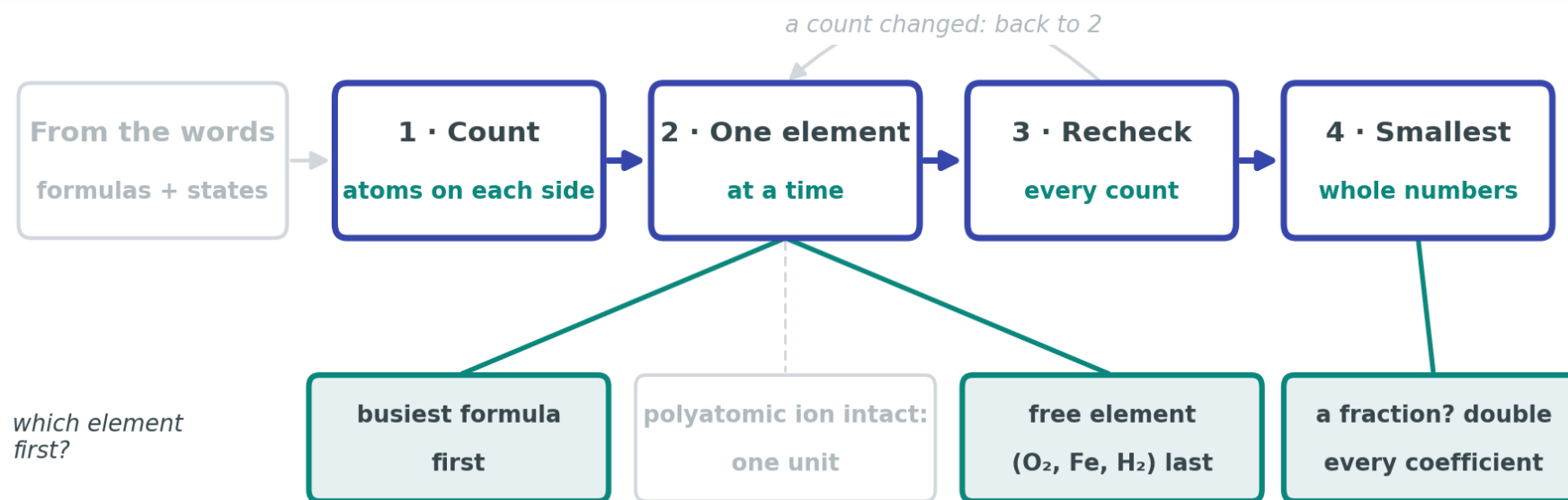
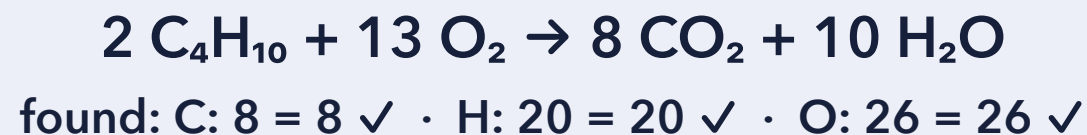
step	equation
C, then H	$\text{C}_4\text{H}_{10} + \text{O}_2 \rightarrow 4 \text{CO}_2 + 5 \text{H}_2\text{O}$
O: right side holds 13	$\text{C}_4\text{H}_{10} + 13/2 \text{O}_2 \rightarrow 4 \text{CO}_2 + 5 \text{H}_2\text{O}$
double it	$2 \text{C}_4\text{H}_{10} + 13 \text{O}_2 \rightarrow 8 \text{CO}_2 + 10 \text{H}_2\text{O}$



C: 8 = 8 ✓ · H: 20 = 20 ✓ · O: 26 = 26 ✓

🧠 **Sense check:** 13 is odd, so no common factor remains. These are already the smallest whole numbers.

✓ Worked example 4: the route



 **Sense check:** C_4H_{10} first, O_2 last. The free element took the odd count as $13/2$, and doubling every coefficient cleared the fraction.



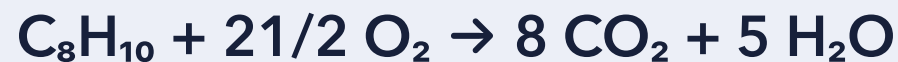
Practice 2: combustion



Xylene, a paint-thinner solvent, burns in air. Balance with smallest whole numbers. What is the sum of **all** the coefficients?

- A. 35
- B. 24.5
- C. 49
- D. 25

✓ Practice 2: answer C



C first: 8 CO₂ · H next: 5 H₂O · O last: 8 × 2 + 5 = 21 on the right, so 21/2 O₂



C: 16 = 16 ✓ · H: 20 = 20 ✓ · O: 42 = 32 + 10 ✓ · sum: 2 + 21 + 16 + 10 = 49 (answer C)

B stopped at the fraction: 1 + 21/2 + 8 + 5 = 24.5, not whole numbers. A doubled only the O₂: 1 + 21 + 8 + 5 = 35, with 42 O on the left against 21 on the right. D rounded 21/2 up to 11: 1 + 11 + 8 + 5 = 25, and 22 O ≠ 21.

 **Sense check:** A fraction clears only when every coefficient doubles. 21 is odd, so 2, 21, 16, 10 share no factor: 49 is the smallest-whole-number sum.

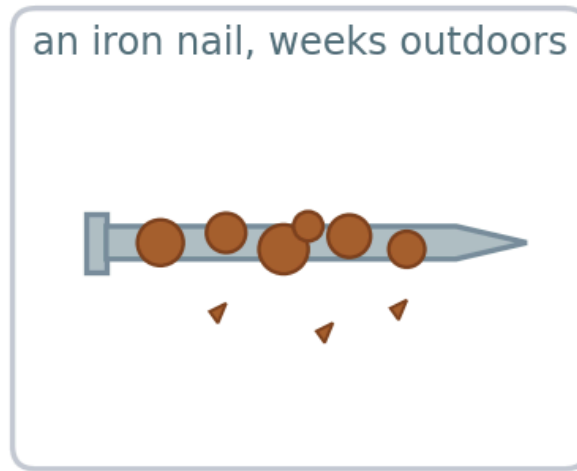
2 · Types of Reactions

Spot the evidence that a reaction occurred, classify any equation as one of the five reaction types, and predict products from the pattern.

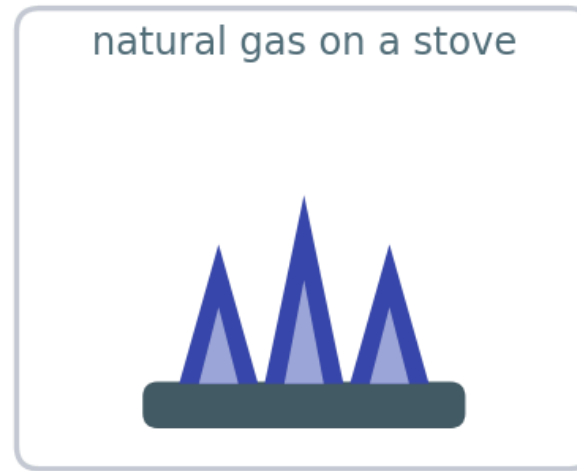
Signs of a new substance



foams up



turns brown and flaky



burns blue

Baking soda and vinegar foam. A nail turns brown and flaky. Stove gas burns blue. Each change makes a new substance, and each leaves a visible sign.

Reactions repeat a few patterns



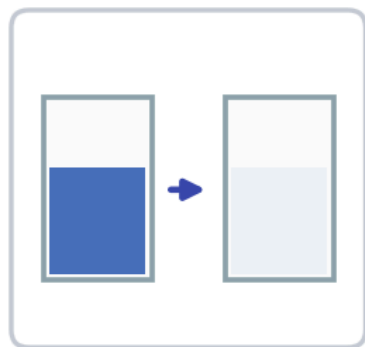
element + element → one compound



element + element → one compound, the same pattern

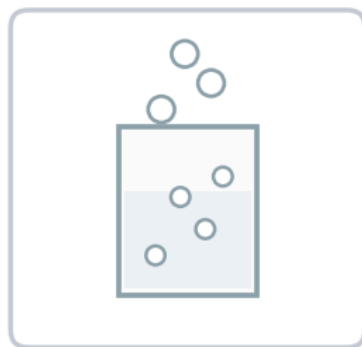
Millions of reactions are known. Nearly all follow a few repeating patterns, visible in the equation's shape. Recognizing the pattern tells what the products must be before any balancing starts.

Evidence a reaction occurred



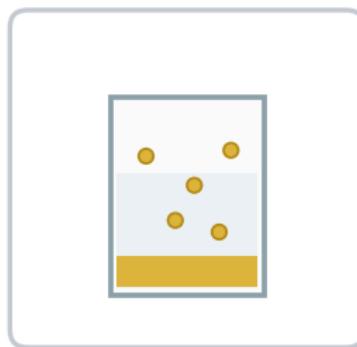
color change

a new color appears



gas forms

bubbles rise out



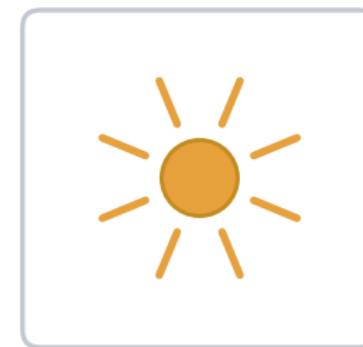
precipitate

a new solid settles out



temperature change

heats up or cools down



light

the mixture glows

New substances have new properties, so the change is visible. Watch for a color change, a gas bubbling out, a solid settling from clear solutions (a **precipitate**), a temperature change, or light.

Reading a pattern



each letter stands for an element or an ion, not one particular substance



$A = \text{Zn}$ · $BC = \text{HCl}$ (B is H, C is Cl) · $AC = \text{ZnCl}_2$ · B leaves as H_2

A pattern is a template: any element or ion can play a letter. Coefficients play no part; the pattern reads only which substances are elements and which are compounds.

The five patterns

combination



decomposition



single displacement



double displacement



combustion



Five shapes suffice: atoms can only join, split apart, or trade partners. The shape carries the classification: how many substances on each side, element or compound. Combustion adds a signature: O_2 consumed, CO_2 and H_2O formed.



The method

1. **Inventory each side.** Count the substances; mark each as element or compound.
2. **Match the shape.** Pick the pattern the inventory fits.
3. **Name the type and complete the products.** If products are missing, the pattern supplies them.



Worked example 1: heating limestone

Step 1 • Inventory each side



one reactant, a compound $\rightarrow$ two products, both compounds • O: 3 = 1 + 2 ✓

Limestone breaks down in a hot kiln. Classify the reaction.

A common first attempt: CO_2 forms, so combustion. Test it.

✓ Worked example 1: solution



one reactant, a compound $\rightarrow$ two products, both compounds

A common first attempt



combustion: fuel + $\text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$ · no O_2 is consumed here ✕

Step 2 · Match the shape

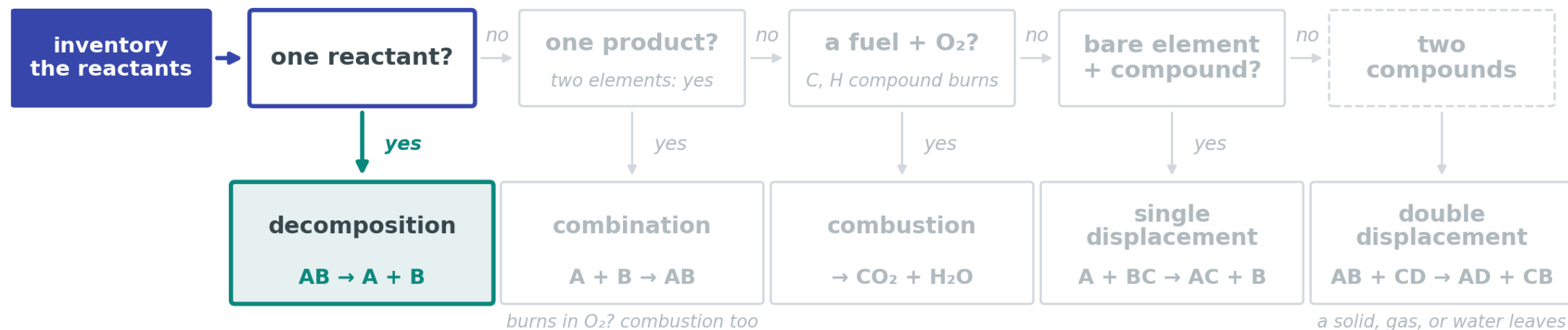
Step 3 · Name the type and complete the products



 **Sense check:** Heat split one compound into two simpler ones. CO_2 appeared without any burning: the products alone cannot name the type.

✓ Worked example 1: the route on the map

$\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$: decomposition
given: one reactant, a compound · found: decomposition



🧠 **Sense check:** The first question settles it: one reactant. The CO₂ never reaches the fuel question. ✓



Worked example 2: propane on a grill

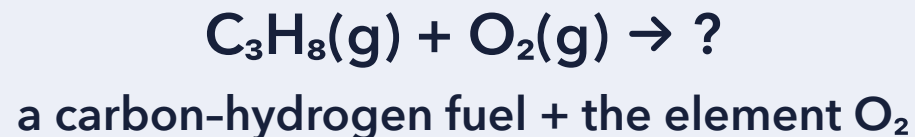
Step 1 • Inventory each side



a carbon-hydrogen fuel + the element O_2 • products not yet written

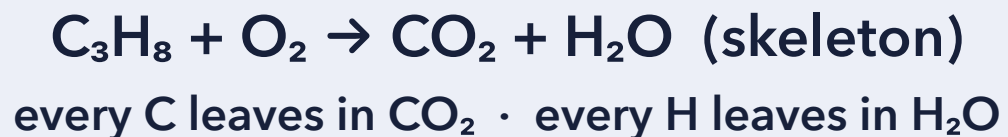
Propane burns in a grill. Classify the reaction, write the products the pattern requires, and balance.

✓ Worked example 2: solution

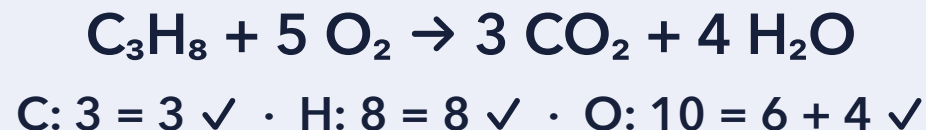


Step 2 • Match the shape

Step 3 • Name the type and complete the products



The balance check

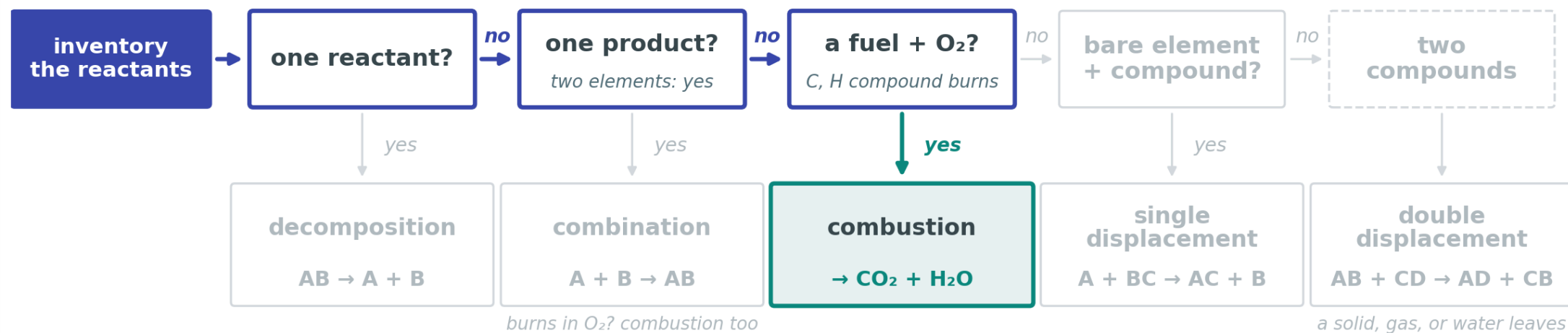


 **Sense check:** Every carbon-hydrogen fuel burns to the same two products. The pattern chose CO_2 and H_2O ; balancing only chose the amounts.

✓ Worked example 2: the route on the map



given: a fuel + O₂ · found: combustion, CO₂ and H₂O



Sense check: Two reactants, and two products, not one: the first two exits pass. A C, H fuel with O₂ exits at combustion, and that exit fixes CO₂ and H₂O. ✓

Your turn: magnesium in acid



bubbles rise as the metal dissolves · H: 2 = 2 ✓ · Cl: 2 = 2 ✓

step	question	answer
1 · Inventory each side	element or compound?	bare element + compound → compound + bare _____
2 · Match the shape	which pattern fits?	A + BC → _____ + B
3 · Name the type		_____ displacement

Complete the three steps.



Mg displaces hydrogen from HCl · the escaping H₂ gas is the visible evidence

⚠ Where this goes wrong

Reading any two-and-two equation as a partner swap. $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$ shows two reactants and two products, but Zn enters as a bare element. Double displacement needs two compounds; a bare element displacing another is single displacement.

Calling every CO_2 producer combustion. $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ releases CO_2 with no O_2 consumed. Combustion consumes O_2 and forms CO_2 and H_2O ; one compound splitting apart is decomposition.

Reading the arrow backwards. Combination builds one product from several reactants; decomposition splits one reactant into several products. $2 \text{HgO} \rightarrow 2 \text{Hg} + \text{O}_2$ starts from a single compound: decomposition.

Calling every bright, hot reaction combustion. $2 \text{Na} + \text{Cl}_2 \rightarrow 2 \text{NaCl}$ gives off heat and light, yet no O_2 is consumed and no CO_2 or H_2O forms. Two elements forming one compound: combination.



Practice 1



A welding torch burns acetylene, C_2H_2 , in pure oxygen. Which type best classifies this reaction?

- A. Double displacement: two reactants form two products, so two pairs traded partners
- B. Combination: the fuel and the oxygen combine into new compounds
- C. Combustion: a carbon-hydrogen fuel consumes O_2 and forms CO_2 and H_2O
- D. Decomposition: the heat breaks the C_2H_2 molecule apart

✓ Practice 1: answer C

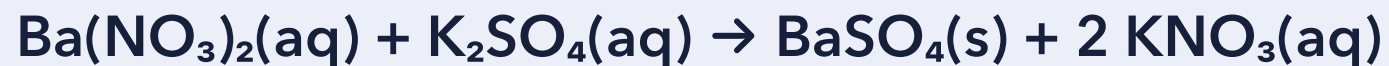
$2 \text{C}_2\text{H}_2 + 5 \text{O}_2 \rightarrow 4 \text{CO}_2 + 2 \text{H}_2\text{O}$: combustion (answer C)
a carbon-hydrogen fuel consumes O_2 and forms CO_2 and H_2O : the full signature

A: a partner swap needs two compounds trading ions, and O_2 is a bare element. B: combination merges everything into one product; two products form here. D: decomposition starts from one reactant; two react here, and the fuel is not falling apart on its own.

 **Sense check:** Fuel and O_2 on the left, CO_2 and H_2O on the right. The same signature classifies every burning hydrocarbon.



Practice 2



two clear solutions mixed; a white solid settles out

Barium nitrate and potassium sulfate solutions are mixed, and a white solid appears.

Which type best classifies this reaction?

- A. Single displacement: barium displaces potassium from its compound
- B. Combination: the two reactants combine into the one solid product
- C. Decomposition: a compound broke down, which is why a solid fell out
- D. Double displacement: two compounds trade partners, and one new pair leaves as a solid

✓ Practice 2: answer D

$\text{Ba}(\text{NO}_3)_2 + \text{K}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2 \text{KNO}_3$: double displacement (answer D)

$\text{AB} + \text{CD} \rightarrow \text{AD} + \text{CB}$ · the new pair BaSO_4 is the precipitate, the visible evidence

A: single displacement needs a bare element among the reactants, and every substance here is a compound. B: combination ends in one product; two form, and KNO_3 stays dissolved. C: decomposition starts from one reactant; two were mixed, and nothing split into simpler substances.

 **Sense check:** Two compounds in, two compounds out, one of them insoluble. The white solid is the evidence that the partners traded.

Two ways to classify the same reaction

sort by shape: combination · decomposition · single displacement · double displacement

how many pieces, element or compound, and how they recombine

sort by driver: precipitation · acid base · redox · combustion

what pushes the reaction forward

$\text{AgNO}_3 + \text{KCl} \rightarrow \text{AgCl} + \text{KNO}_3$ · shape: double displacement · driver: precipitation

$\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$ · shape: single displacement · driver: redox

One reaction, a name from each scheme. Double displacement splits by driver into precipitation and neutralization; single displacement and combustion are both redox.

Practice 6: both lenses



a clear solution · a gray-white solid collects on the ribbon

Magnesium ribbon stands in silver nitrate solution, and a solid builds up on it. Classify this reaction once per scheme. Which pair of names is correct?

- A. Double displacement, and a precipitation
- B. Single displacement, and a precipitation
- C. Double displacement, and a redox
- D. Single displacement, and a redox

✓ Practice 6: answer D

$\text{Mg} + 2 \text{AgNO}_3 \rightarrow \text{Mg}(\text{NO}_3)_2 + 2 \text{Ag}$: single displacement, and a redox (answer D)
bare element + compound $\rightarrow$ compound + bare element · Mg: 1 = 1 ✓ · Ag: 2 = 2 ✓ · N: 2 = 2 ✓ ·
O: 6 = 6 ✓

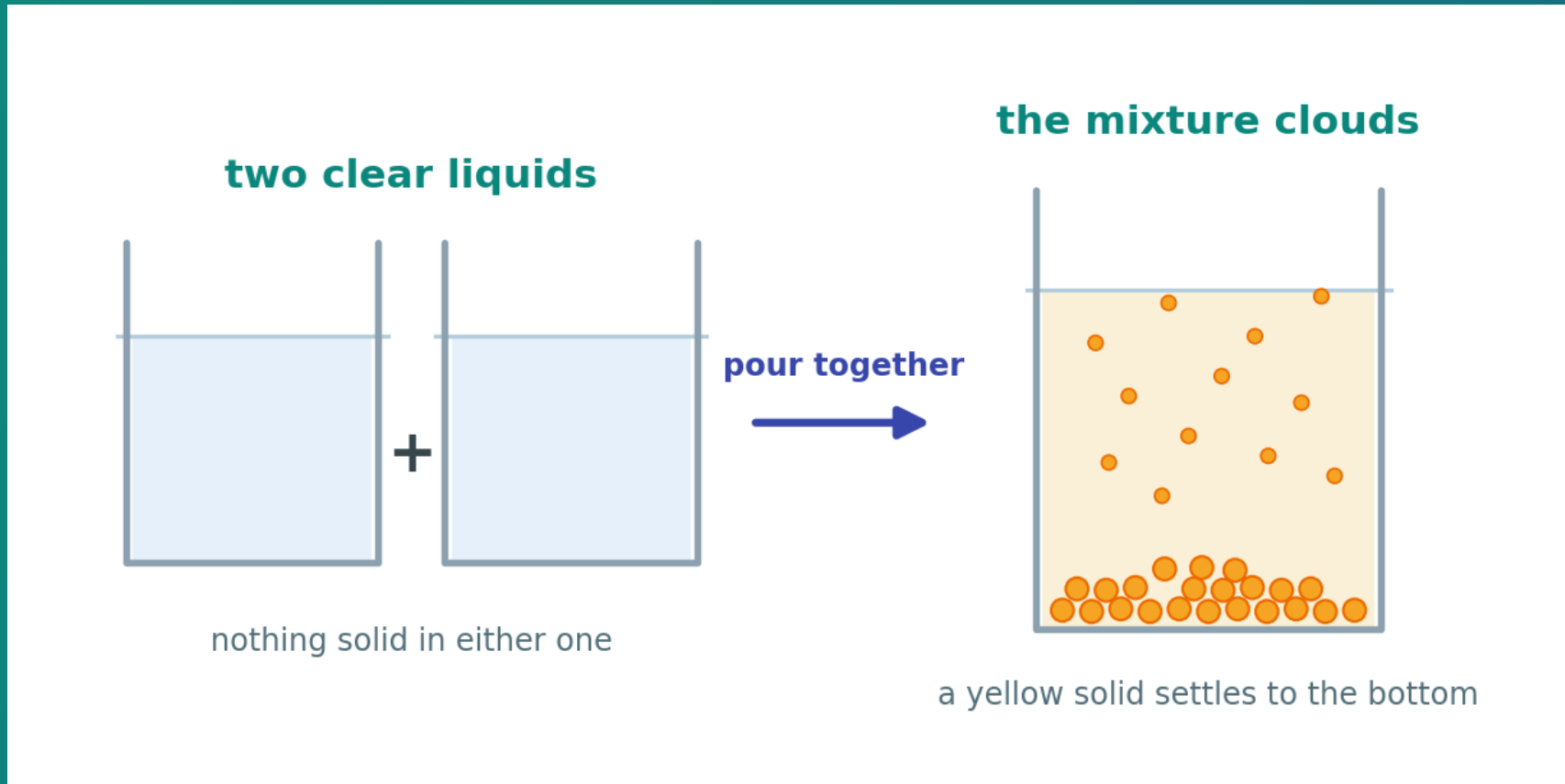
A took both names from the surface: a solid appeared, and two substances went in and two came out. B has the shape right, but the solid is silver metal pushed out of its compound, not a new ion pair leaving the water; precipitation belongs to double displacement. C has the driver right, but Mg enters as a bare element, and a partner swap needs two compounds. Single displacement is driven by electron transfer: a redox.

 **Sense check:** A solid appearing is not proof of a precipitate. Ask what the solid is: a new ion pairing, or an element pushed out of its compound.

3 · Solubility Rules & Precipitation

Predict whether mixing two solutions makes a precipitate: swap the partners, check each new compound against the solubility rules, and write the balanced equation with states.

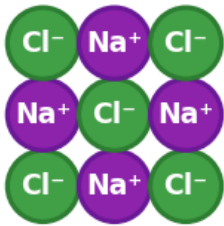
A solid from two clear liquids



Two beakers, two clear liquids. Poured together, they instantly cloud, and a bright yellow solid settles to the bottom. No solid went in.

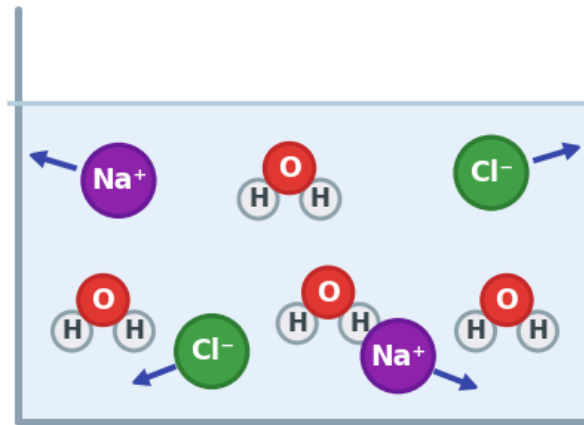
What dissolving looks like

dissolving an ionic solid



NaCl(s)

every ion held by its neighbors



NaCl(aq)

separated ions, each on its own path

the ions separate and travel independently

Dissolving an ionic solid means the ions separate and travel independently. The solid's Na^+ and Cl^- leave the stack one ion at a time and spread through the water.

Soluble or insoluble

NaCl(aq) is soluble: it dissolves, and its ions spread through the water

(aq) = aqueous, dissolved in water

AgCl(s) is insoluble: it stays a solid

(s) = solid; it does not dissolve

An ionic compound in water dissolves or it does not. Dissolving is a tug-of-war: water pulls the ions apart, and for a few pairings the ion-ion attraction wins, so the compound stays solid.

Check the always-soluble ions first

group 1 cations (Li^+ , Na^+ , K^+ , ...) · NH_4^+ · NO_3^-
compounds of these never precipitate · any new pairing holding one stays (aq)

Three families never precipitate. Check a new pairing against this short list first; it settles most cases. Only a pairing with none of the three needs the chart's exception lines.

The solubility rules

SOLUBILITY RULES

ionic compounds in water: one line decides each compound

SOLUBLE: dissolves → (aq)

group 1 cations · NH_4^+

Li^+ Na^+ K^+ ... always soluble, no exceptions

nitrates NO_3^- · acetates $\text{C}_2\text{H}_3\text{O}_2^-$

always soluble, no exceptions

chlorides · bromides · iodides

Cl^- Br^- I^- : soluble

EXCEPT with

Ag^+

Pb^{2+}

Hg_2^{2+}

sulfates SO_4^{2-}

soluble

EXCEPT with

Ba^{2+}

Sr^{2+}

Pb^{2+}

INSOLUBLE: solid → (s)

carbonates CO_3^{2-}

insoluble

EXCEPT with

group 1

NH_4^+

sulfides S^{2-}

insoluble

EXCEPT with

group 1

NH_4^+

hydroxides OH^-

insoluble

EXCEPT with

group 1

Ba^{2+}

an EXCEPT partner reverses the rule:



red: forms the solid (s)



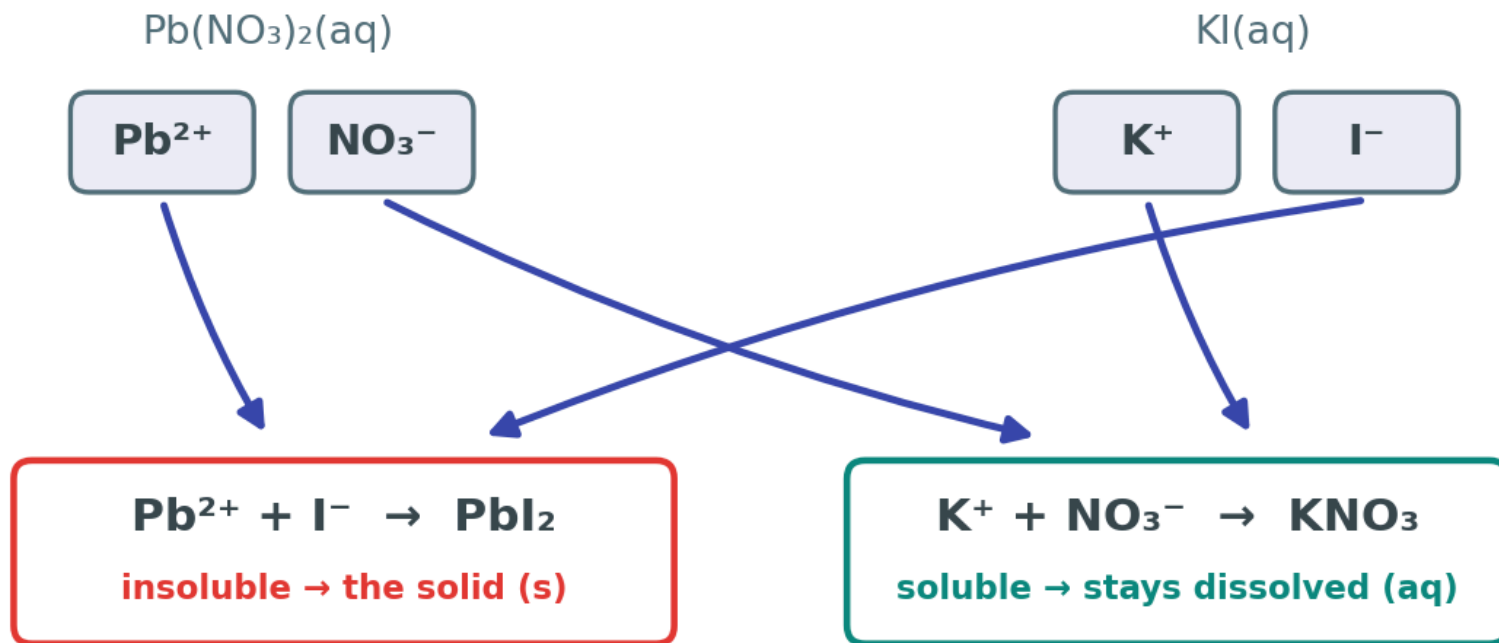
teal: stays dissolved (aq)

NAGSAG: Nitrates · Acetates · Group 1 · Sulfates · Ammonium · Group 17 halides
the mostly-soluble families · only the sulfates and the halides (Cl^- Br^- I^-) carry exception lists

One line decides each compound: read its ions, find their line, apply the exception list.

Mixing two solutions: the partners swap

double replacement: the partners swap



new pairings: each checked against the solubility rules

Dissolved compounds are separated ions. Mixing two solutions lets each cation meet the other anion. A new pairing the rules call insoluble forms a solid: a **precipitate**.

Warm-up: formulas from charge balance



wanted: the neutral formula each ion pair builds

ions	charge balance	formula
$\text{Ba}^{2+} + \text{OH}^{-}$	$1(2+) + 2(1-) = 0$	$\text{Ba}(\text{OH})\underline{\hspace{1cm}}$
$\text{Al}^{3+} + \text{SO}_4^{2-}$	$2(3+) + 3(2-) = 0$	$\underline{\hspace{1cm}}$

Complete both formulas.



$1(2+) + 2(1-) = 0 \checkmark \cdot 2(3+) + 3(2-) = 0 \checkmark$ · charge balance sets the subscripts, exactly as when naming ionic compounds



The method

1. **Swap the partners.** New cation-anion pairs.
2. **Build each new formula.** Charge balance sets subscripts.
3. **Check each product against the rules.** Soluble $\rightarrow$ (aq); insoluble $\rightarrow$ (s).
4. **Write the balanced equation with states.** Both products (aq): no reaction.



Worked example 1: soluble or insoluble



wanted: the state of each in water, (aq) or (s)

White marble is CaCO_3 . Potash fertilizer, K_2CO_3 , is spread as a clear solution. Assign each compound its state in water.

✓ Worked example 1: solution



wanted: the state of each in water, (aq) or (s)

Read the ions



both compounds are carbonates: the carbonate line decides both

Apply the carbonate line

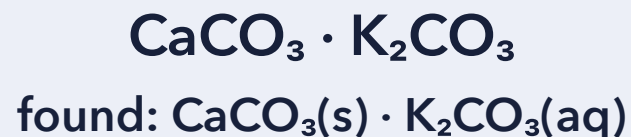
Carbonates are insoluble except with group 1 cations or NH_4^+ .



Ca^{2+} : not on the exception list $\rightarrow$ solid $\cdot$ K^+ : group 1 $\rightarrow$ the exception applies, dissolved

 **Sense check:** Same anion, opposite states. A rule reads the pair of ions; the exception list is part of the rule.

✓ Worked example 1: the route on the chart



SOLUBLE: dissolves → (aq)	INSOLUBLE: solid → (s)
group 1 cations · NH_4^+ $\text{Li}^+ \text{Na}^+ \text{K}^+ \dots$: always soluble	carbonates CO_3^{2-} insoluble EXCEPT with group 1 NH_4^+ $\text{CaCO}_3 \rightarrow (\text{s})$ $\text{K}_2\text{CO}_3 \rightarrow (\text{aq})$
nitrates NO_3^- · acetates $\text{C}_2\text{H}_3\text{O}_2^-$ always soluble, no exceptions	sulfides S^{2-} insoluble EXCEPT with group 1 NH_4^+
$\text{Cl}^- \cdot \text{Br}^- \cdot \text{I}^-$ chlorides, bromides, iodides: soluble EXCEPT with Ag^+ Pb^{2+} Hg_2^{2+}	hydroxides OH^- insoluble EXCEPT with group 1 Ba^{2+}
sulfates SO_4^{2-} soluble EXCEPT with Ba^{2+} Sr^{2+} Pb^{2+}	

 **Sense check:** Both compounds read the carbonate line: insoluble. K^+ is group 1, an EXCEPT partner, so K_2CO_3 flips to (aq). Ca^{2+} is not on the list: $\text{CaCO}_3(\text{s})$.



Worked example 2: $\text{AgNO}_3 + \text{NaCl}$

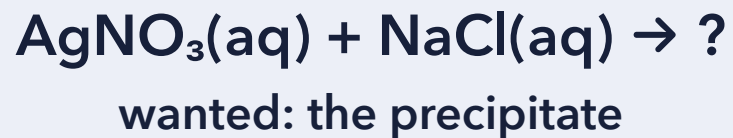


wanted: the precipitate

Silver nitrate solution meets salt solution in photographic processing. Both are clear and colorless; mixing them turns the beaker milky white. Predict the solid.

A common first attempt: check AgNO_3 and NaCl against the rules: both soluble, so no solid should form. Test it.

✓ Worked example 2: testing the first attempt



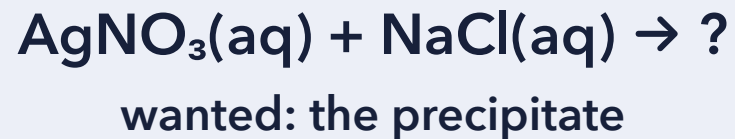
A common first attempt

AgNO_3 soluble ✓ · NaCl soluble ✓ → no solid?
 nitrates: always soluble · group 1: always soluble, yet the mixture turns white ✕

Both reactants pass the rules. That is what (aq) already records: they arrived dissolved. The white solid must be a compound neither beaker held.

 **Sense check:** The mixed beaker holds four ions moving independently: Ag^+ , NO_3^- , Na^+ , Cl^- . Two of them meet here for the first time.

✓ Worked example 2: the swap



Step 1 · Swap the partners

Step 2 · Build each new formula

Ag^+ pairs with Cl^- ; Na^+ pairs with NO_3^- . Both new pairs balance one-to-one.

new pairs: AgCl and NaNO_3



 **Sense check:** Two dissolved compounds went in; two new pairings remain to be checked against the rules.

✓ Worked example 2: states and the equation



new pairs built · wanted: each product's state

Step 3 · Check each product against the rules



chlorides: soluble except Ag^+ , Pb^{2+} , Hg_2^{2+} ; Ag^+ is on the list · nitrates: always soluble

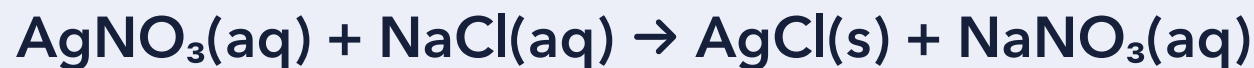
Step 4 · Write the balanced equation with states



Ag: 1 = 1 · NO₃: 1 = 1 · Na: 1 = 1 · Cl: 1 = 1 · balanced as written

 **Sense check:** The milky white is AgCl, the one new pairing the rules call insoluble. Na⁺ and NO₃⁻ never left the water.

✓ Worked example 2: the route on the chart

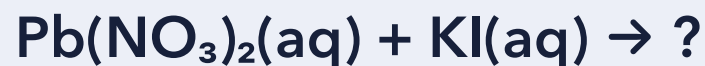


found: $\text{AgCl}(\text{s}) \cdot \text{NaNO}_3(\text{aq})$

SOLUBLE: dissolves → (aq)	INSOLUBLE: solid → (s)
group 1 cations · NH_4^+ Li ⁺ Na ⁺ K ⁺ ...: always soluble	carbonates CO_3^{2-} insoluble EXCEPT with group 1 NH_4^+
nitrates NO_3^- · acetates $\text{C}_2\text{H}_3\text{O}_2^-$ $\text{NaNO}_3 \rightarrow (\text{aq})$ always soluble, no exceptions	sulfides S^{2-} insoluble EXCEPT with group 1 NH_4^+
$\text{Cl}^- \cdot \text{Br}^- \cdot \text{I}^-$ $\text{AgCl} \rightarrow (\text{s})$ chlorides, bromides, iodides: soluble EXCEPT with Ag^+ Pb^{2+} Hg_2^{2+}	hydroxides OH^- insoluble EXCEPT with group 1 Ba^{2+}
sulfates SO_4^{2-} soluble EXCEPT with Ba^{2+} Sr^{2+} Pb^{2+}	

 **Sense check:** NaNO_3 reads the nitrate line: always soluble. AgCl reads the chloride line, and Ag^+ is an EXCEPT partner: the verdict flips to (s).

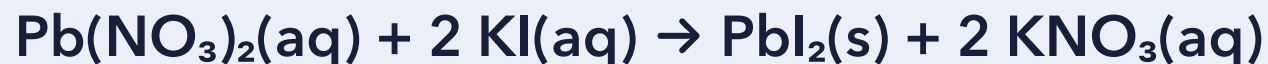
Your turn: $\text{Pb}(\text{NO}_3)_2 + \text{KI}$



two clear solutions · mixing makes a bright yellow solid

step	question	answer
1 · swap the partners	which new pairs form?	Pb^{2+} with I^- · K^+ with NO_3^-
2 · build each new formula	$1(2+) + 2(1-) = 0$	$\text{PbI}_{\underline{\hspace{1cm}}}$ and KNO_3
3 · check each product	which lines of the rules?	$\text{PbI}_2 \rightarrow (\underline{\hspace{1cm}})$ · $\text{KNO}_3 \rightarrow (\underline{\hspace{1cm}})$
4 · write the balanced equation	coefficients	$\underline{\hspace{1cm}}$

Complete the prediction.



PbI_2 : $1(2+) + 2(1-) = 0 \checkmark$ · iodides: soluble except Ag^+ , Pb^{2+} , Hg_2^{2+} · nitrates: always soluble · PbI_2 settles brilliant yellow

⚠ Where this goes wrong

Checking the rules on the starting solutions. AgNO_3 and NaCl both pass: they were already dissolved. The rules judge the products of the swap: Ag^+ with Cl^- gives AgCl , and the chloride line puts Ag^+ on its exception list.

Calling the soluble product the precipitate. $\text{BaCl}_2 + \text{Na}_2\text{SO}_4$ gives two new pairs. NaCl passes the rules, so its ions stay dissolved. The solid is the pairing that does not: BaSO_4 .

Expecting a starting compound to fall out. A compound that arrived dissolved stays dissolved. The precipitate is always a new pairing: ions that arrived in different beakers.

Dragging old subscripts into new formulas. $\text{Pb}(\text{NO}_3)_2 + \text{KI}$: copying KI 's one-to-one ratio gives PbI , and $1(+2) + 1(-1) = +1$, not neutral. Charge balance builds the new formula: PbI_2 , $1(+2) + 2(-1) = 0$.



Practice 1



two clear solutions are mixed

Silver nitrate solution is poured into sodium sulfide solution. Which precipitate forms, if any?

- A. Ag_2S
- B. NaNO_3
- C. AgS
- D. No precipitate: both new pairings stay dissolved

✓ Practice 1: answer A



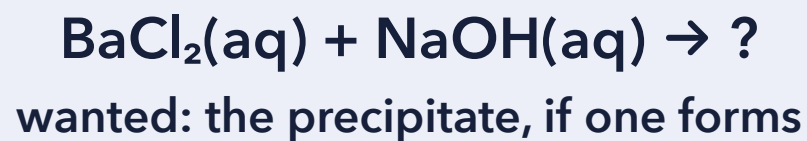
sulfides: insoluble except group 1 and NH_4^+ · Ag_2S : $2(1+) + 1(2-) = 0$ ✓ · Ag: $2 = 2$ · S: $1 = 1$ · Na: $2 = 2$ · NO_3 : $2 = 2$

B is the soluble product: sodium is group 1 and nitrate is always soluble, so NaNO_3 's ions stay dissolved. C has the right pairing and the wrong formula: AgS gives $1(1+) + 1(2-) = -1$, so charge balance needs two Ag^+ per S^{2-} . D skipped the sulfide line: Ag^+ is not group 1 or NH_4^+ , so Ag_2S falls out.

 **Sense check:** Of the four ions mixed, only $\text{Ag}^+ + \text{S}^{2-}$ lands on an insoluble line. One insoluble pairing, one solid: black Ag_2S , the same compound that tarnishes silverware.



Worked example 3: $\text{BaCl}_2 + \text{NaOH}$



Solutions of barium chloride and sodium hydroxide are mixed. Work all four steps.

✓ Worked example 3: the swap

$\text{BaCl}_2(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow ?$
 wanted: the precipitate, if one forms

Step 1 · Swap the partners

Step 2 · Build each new formula

Ba^{2+} pairs with OH^- ; Na^+ pairs with Cl^- . Ba^{2+} needs two OH^- to reach zero charge.

new pairs: $\text{Ba}(\text{OH})_2$ and NaCl

$\text{Ba}(\text{OH})_2$: $1(+2) + 2(-1) = 0 \checkmark$ · NaCl : $1(+1) + 1(-1) = 0 \checkmark$

 **Sense check:** The new formulas come from charge balance, never from the reactants' subscripts.

✓ Worked example 3: solution



new pairs built · wanted: each product's state

Step 3 · Check each product against the rules



hydroxides: insoluble except group 1 and Ba^{2+} · group 1: always soluble

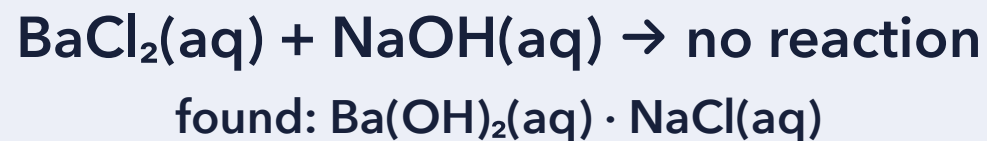
Step 4 · Write the balanced equation with states



both new pairings stay dissolved: nothing leaves the solution

 **Sense check:** The mixed beaker stays clear: it holds four kinds of dissolved ions. A precipitate needs one insoluble pairing, and here there is none.

✓ Worked example 3: the route on the chart



SOLUBLE: dissolves → (aq)	INSOLUBLE: solid → (s)
group 1 cations · NH_4^+ $\text{Li}^+ \text{Na}^+ \text{K}^+ \dots$: always soluble NaCl → (aq)	carbonates CO_3^{2-} insoluble EXCEPT with group 1 NH_4^+
nitrates NO_3^- · acetates $\text{C}_2\text{H}_3\text{O}_2^-$ always soluble, no exceptions	sulfides S^{2-} insoluble EXCEPT with group 1 NH_4^+
$\text{Cl}^- \cdot \text{Br}^- \cdot \text{I}^-$ chlorides, bromides, iodides: soluble EXCEPT with Ag^+ Pb^{2+} Hg_2^{2+}	hydroxides OH^- insoluble EXCEPT with group 1 Ba^{2+} $\text{Ba}(\text{OH})_2 \rightarrow (\text{aq})$
sulfates SO_4^{2-} soluble EXCEPT with Ba^{2+} Sr^{2+} Pb^{2+}	

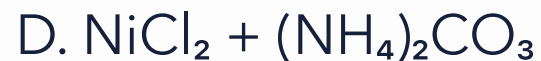
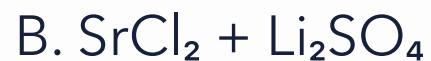
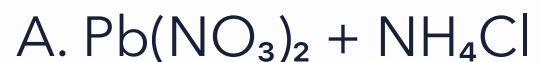
 **Sense check:** NaCl reads the group 1 line: always soluble. Ba^{2+} is an EXCEPT partner on the hydroxide line, so $\text{Ba}(\text{OH})_2$ flips to (aq). Two teal verdicts: nothing leaves the solution.



Practice 2

four beakers · one pair of clear solutions mixed in each
three beakers turn cloudy · wanted: the mixture that stays clear

Which mixture stays clear?



✓ Practice 2: answer C

new pairs: $\text{BaI}_2 \rightarrow (\text{aq}) \cdot \text{NaNO}_3 \rightarrow (\text{aq}) \rightarrow \text{no reaction (answer C)}$

BaI_2 : $1(+2) + 2(-1) = 0$ ✓ · iodides: soluble except Ag^+ , Pb^{2+} , Hg_2^{2+} · Ba^{2+} sits on the sulfate list, not the iodide list

A: $\text{PbCl}_2(\text{s})$ · B: $\text{SrSO}_4(\text{s})$ · D: $\text{NiCO}_3(\text{s})$

Pb^{2+} : a chloride exception · Sr^{2+} : a sulfate exception · Ni^{2+} : not group 1, not NH_4^+ · each partner (NH_4NO_3 , LiCl , NH_4Cl) stays (aq)

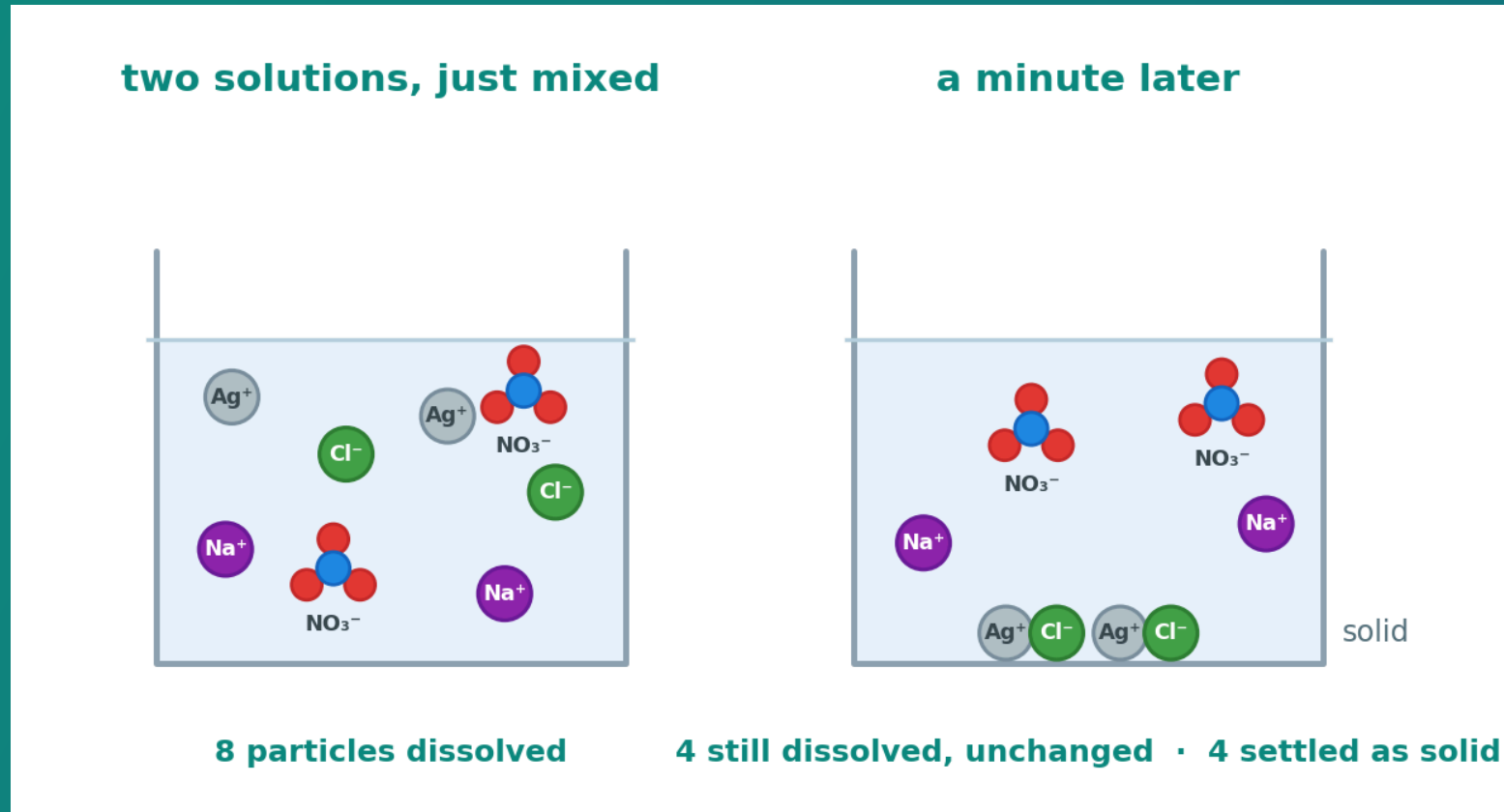
A trusted the always-soluble ions: NH_4^+ and NO_3^- stay dissolved, but the new pair PbCl_2 is on the chloride exception list. B remembered only barium on the sulfate list; Sr^{2+} is there too. D applied the ammonium exception to the wrong pairing: $(\text{NH}_4)_2\text{CO}_3$ arrived dissolved, but NiCO_3 has no exception.

 **Sense check:** Barium fails the sulfate line and passes the halide lines. An exception list belongs to one anion: no cation precipitates with everything.

4 • Net Ionic Equations

Write the molecular, complete ionic, and net ionic equations for a reaction in solution, cancel the spectator ions, and check that the result balances in atoms and charge.

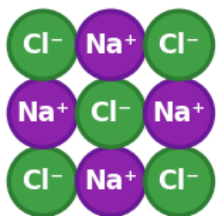
The solid takes only two kinds of particles



Two clear solutions are mixed; a white solid settles. Most dissolved particles are still floating afterward, unchanged. Only two kinds left the water.

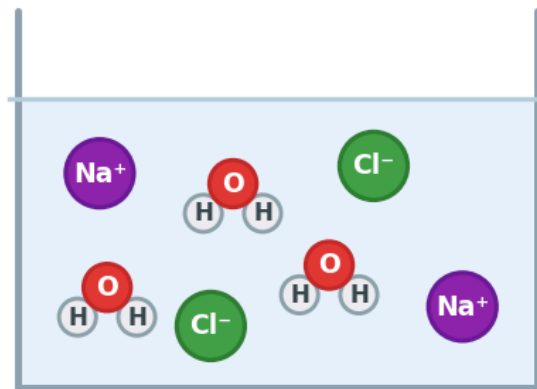
Dissolved means separated into ions

dissolving separates the ions



NaCl(s)

ions locked in a grid



NaCl(aq)

separated ions, each moving on its own

the water holds $\text{Na}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$: no NaCl particle survives

A soluble ionic compound does not dissolve as molecules. It exists in the water as **separated ions**. Writing the ions shows which of them react and which never change.

Three views of the same reaction

molecular



write each (aq) compound as its ions; AgCl(s) stays intact

complete ionic



cancel the ions that match on both sides: Na⁺, NO₃⁻

net ionic



one beaker, one reaction: three ways to write it

The molecular equation lists whole compounds. The two ionic views show what the water actually holds.



Sense check: "Molecular equation" means written as whole formulas; most compounds in these equations are ionic.

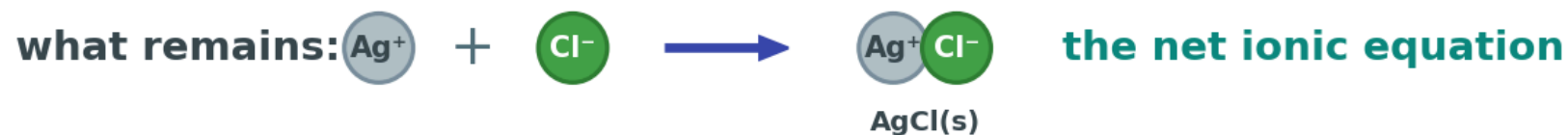
Spectator ions

dissolved before, dissolved after: the spectators cancel



Na^+ and NO_3^- : identical on both sides, unchanged

still dissolved, still spectators



charge check: $(+1) + (-1) = 0$ on the left · the solid is 0 on the right ✓

A **spectator ion** appears identical on both sides. Cancel the spectators; what remains is the **net ionic equation**. Any soluble silver salt plus any soluble chloride gives this same net line.

What separates on the page

written as separated ions: only solutes that separate completely in water
the soluble ionic compounds · the few acids that ionize fully, the strong acids

kept whole: everything else
acids that barely ionize: $\text{HC}_2\text{H}_3\text{O}_2(\text{aq})$ stays whole · insoluble solids: $\text{AgCl}(\text{s})$ stays intact · covalent molecules: $\text{H}_2\text{O}(\text{l})$ and sugar

A formula earns ion notation only by existing in the water fully separated. In these reactions that means the soluble salts and the strong acids; everything else keeps its formula whole.



The method

1. **Write the molecular equation with states.**
2. **Write each (aq) compound as its ions.** Keep charges and coefficients; polyatomic ions stay whole; (s) stays intact.
3. **Cancel the spectator ions.**
4. **Check atoms and charge.** Both must balance.



Worked example 1: silver chloride

Step 1 • Write the molecular equation with states



states from the solubility rules: AgCl insoluble, the rest soluble

A drop of silver nitrate turns salty water cloudy white, a standard test for chloride. Write the complete ionic equation, then the net ionic equation.

✓ Worked example 1: the complete ionic equation



molecular · states from the solubility rules

Step 2 · Write each (aq) compound as its ions



complete ionic · charge: left $1(+1) + 1(-1) + 1(+1) + 1(-1) = 0$ · right $0 + 1(+1) + 1(-1) = 0$ ✓

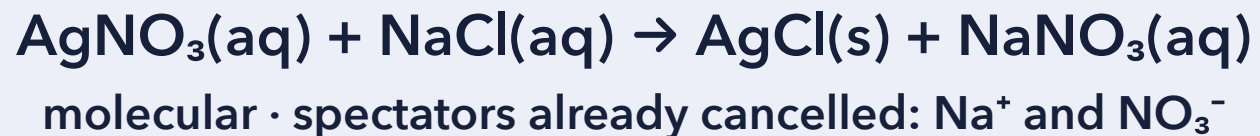
Step 3 · Cancel the spectator ions



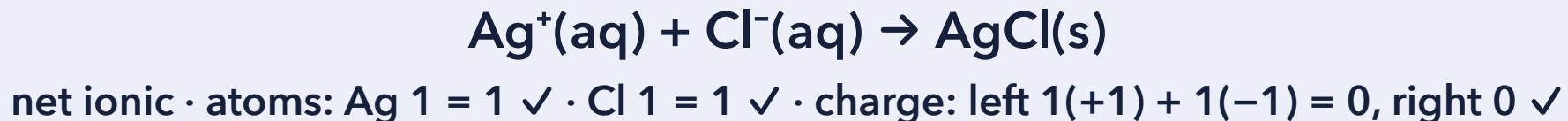
spectators: Na^+ and NO_3^-

 **Sense check:** The compound formulas are gone; the water's actual contents are on the page, and the two ions that never react are struck out.

✓ Worked example 1: the net ionic equation



Step 4 · Check atoms and charge

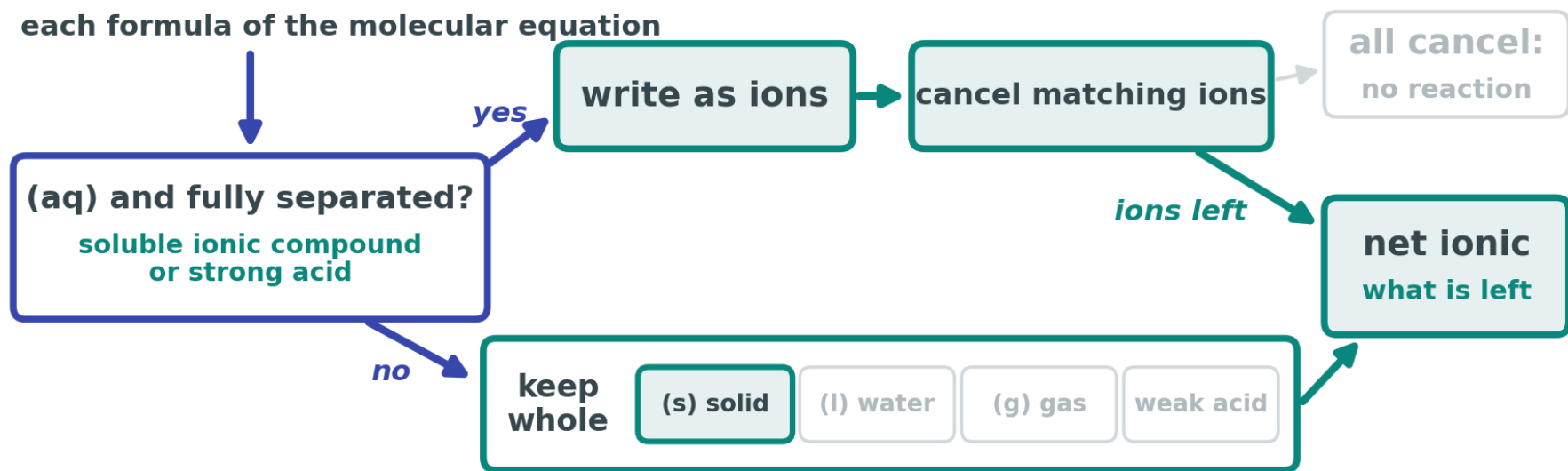


🧠 **Sense check:** The net ionic equation balances twice: every atom matches, and both sides carry zero total charge.

✓ Worked example 1: the route on the strip



split: AgNO_3 , NaCl , NaNO_3 · kept whole: $\text{AgCl}(\text{s})$ · found: $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$

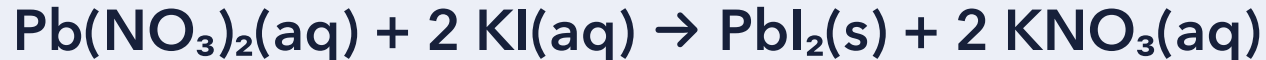


Sense check: Three formulas took the yes branch; the solid took no. Na^+ and NO_3^- cancel, and the ions that build AgCl are all that is left. ✓



Worked example 2: lead iodide

Step 1 • Write the molecular equation with states

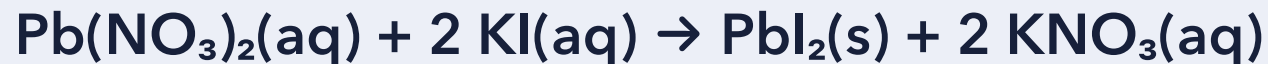


PbI_2 insoluble: the yellow solid of the golden-rain demonstration

Write the complete ionic equation, then the net ionic equation.

A common first attempt for the net: $\text{Pb}^{2+}(\text{aq}) + \text{I}^{-}(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$. Test it.

✓ Worked example 2: the complete ionic equation



molecular · PbI_2 is the solid

Step 2 · Write each (aq) compound as its ions



charge: left $1(+2) + 2(-1) + 2(+1) + 2(-1) = 0$ · right $0 + 2(+1) + 2(-1) = 0$ ✓

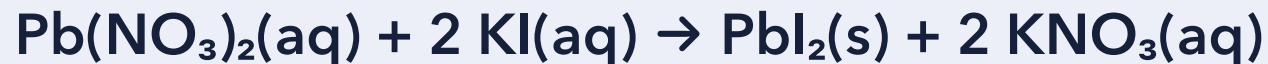
Step 3 · Cancel the spectator ions



spectators: K^{+} and NO_3^{-} · Pb^{2+} and 2I^{-} have no match to cancel

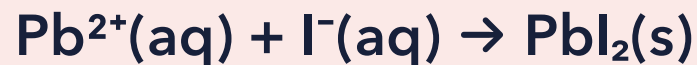
 **Sense check:** Nitrate separates and cancels as one whole unit, never as N and O pieces, and its coefficient 2 stays with it.

✓ Worked example 2: the net ionic equation



molecular · spectators already cancelled: K^+ and NO_3^-

A common first attempt



atoms: I 1 $\neq$ 2 ✗ · charge: left 1(+2) + 1(−1) = +1, right 0 ✗

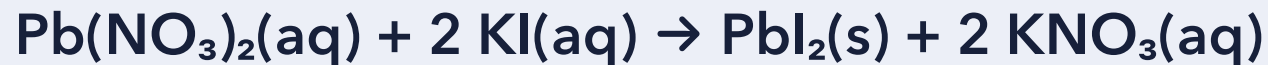
Step 4 · Check atoms and charge



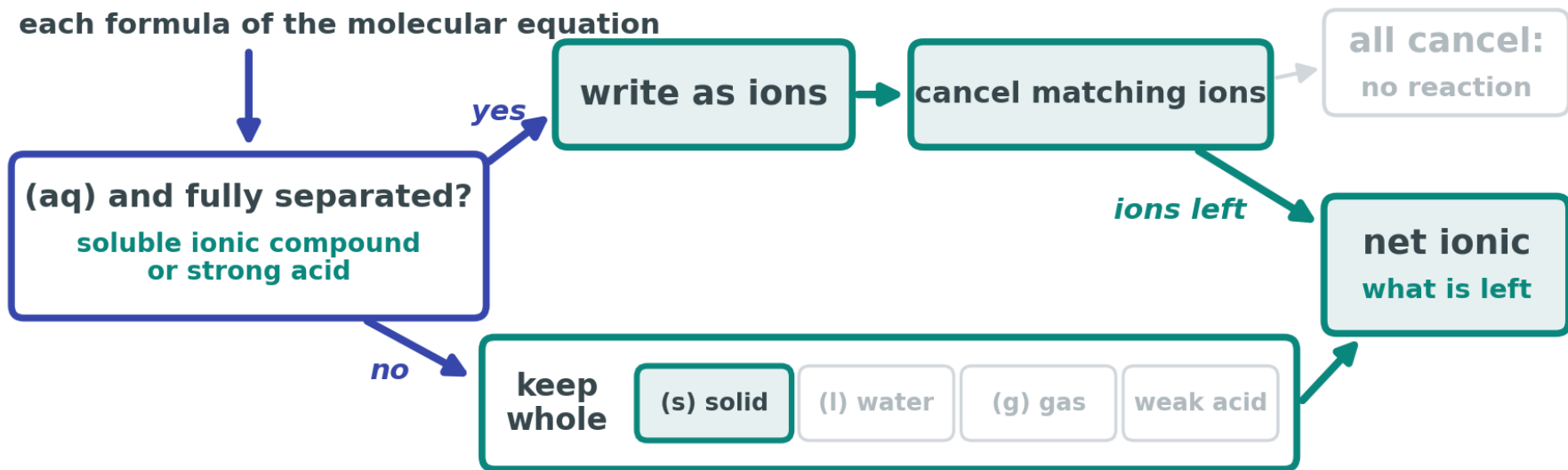
atoms: Pb 1 = 1 ✓ · I 2 = 2 ✓ · charge: left 1(+2) + 2(−1) = 0, right 0 ✓

 **Sense check:** The solid is neutral, so the ions that build it must sum to zero. Keeping the coefficients is what makes both checks pass.

✓ Worked example 2: the route on the strip



split: $\text{Pb}(\text{NO}_3)_2$, KI, KNO_3 · kept whole: $\text{PbI}_2(\text{s})$ · found: $\text{Pb}^{2+}(\text{aq}) + 2 \text{I}^{-}(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$



Sense check: Same route as silver chloride: the solid took the no branch. K^{+} and NO_3^{-} cancel; Pb^{2+} and 2I^{-} are left, with their coefficient. ✓

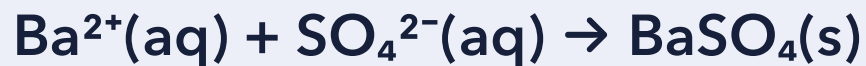
Your turn: barium sulfate



Step 1 done · BaSO_4 insoluble: the X-ray contrast a patient drinks

step	your work
2 · write each (aq) compound as its ions	$\text{Ba}^{2+} + 2 \text{Cl}^- + 2 \text{Na}^+ + \underline{\hspace{1cm}} \rightarrow \text{BaSO}_4(\text{s}) + 2 \text{Na}^+ + 2 \text{Cl}^-$
3 · cancel the spectator ions	$\underline{\hspace{1cm}}$ and $\underline{\hspace{1cm}}$
4 · check atoms and charge	net: $\underline{\hspace{1cm}}$ · left charge $\underline{\hspace{1cm}}$ = right charge 0

Complete steps 2 through 4.



spectators: Na^+ , Cl^- · atoms: Ba 1 = 1 ✓ · S 1 = 1 ✓ · O 4 = 4 ✓ · charge: left $1(+2) + 1(-2) = 0$, right 0 ✓

 **Sense check:** Sulfate stayed whole from the first line to the last. Its charge comes from the memorized list, and the charge check depends on it.

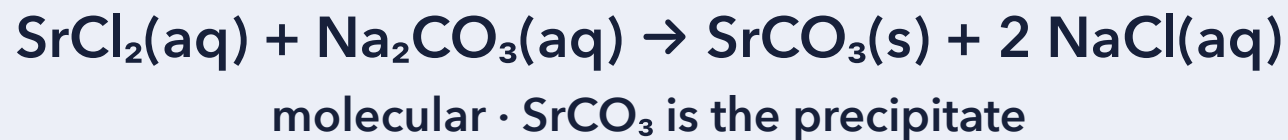
⚠ Where this goes wrong

Splitting the solid into ions. Writing AgCl(s) as $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq})$ lets every ion cancel, and the equation claims nothing happened. A solid formed. Only (aq) compounds separate; (s) stays intact.

Breaking up a polyatomic ion. Dissolved nitrate is $\text{NO}_3^-(\text{aq})$, one whole unit with one charge. Cancel nitrate as nitrate, never as separate N and O.

Dropping the charges. $\text{Ag(aq)} + \text{Cl(aq)} \rightarrow \text{AgCl(s)}$ shows neutral atoms the water does not contain. Without charges, the charge check cannot be run.

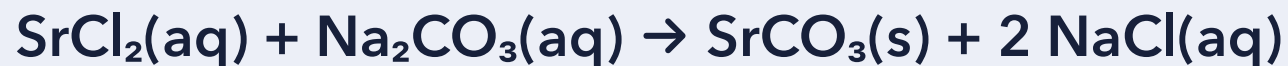
Stopping at the complete ionic equation. If K^+ and NO_3^- still stand on both sides, nothing has been cancelled. The net ionic equation keeps only the ions that build the solid.

 Practice 1

Strontium salts color fireworks red. Aqueous SrCl_2 and Na_2CO_3 are mixed, and SrCO_3 precipitates. Which is the correct net ionic equation?

- A. $\text{SrCl}_2(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{SrCO}_3(\text{s}) + 2 \text{NaCl}(\text{aq})$
- B. $\text{Sr}^{2+}(\text{aq}) + 2 \text{Cl}^{-}(\text{aq}) + 2 \text{Na}^{+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{SrCO}_3(\text{s}) + 2 \text{Na}^{+}(\text{aq}) + 2 \text{Cl}^{-}(\text{aq})$
- C. $\text{Sr}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{SrCO}_3(\text{s})$
- D. $\text{Sr}^{+}(\text{aq}) + \text{CO}_3^{-}(\text{aq}) \rightarrow \text{SrCO}_3(\text{s})$

✓ Practice 1: answer C



the molecular equation, states from the solubility rules



atoms: Sr 1 = 1 ✓ · C 1 = 1 ✓ · O 3 = 3 ✓ · charge: left 1(+2) + 1(−2) = 0, right 0 ✓

A is the molecular equation; nothing has been written as ions. B is the complete ionic equation: Na^+ and Cl^- still stand on both sides, uncanceled. D halves both charges: strontium is a group 2 metal, Sr^{2+} , and carbonate is CO_3^{2-} , so its $1(+1) + 1(-1) = 0$ only looks balanced.

 **Sense check:** An equation can pass the charge check with two wrong charges. Assign each ion's real charge first, then check.

Worked example 3: sodium chloride and potassium nitrate

Step 1 • Write the molecular equation with states



partners swapped; the solubility rules mark every compound (aq): no solid

The two solutions are mixed and stay clear: no solid, no gas, no color change. Write the complete ionic equation, then the net ionic equation.

✓ Worked example 3: solution



molecular · every compound soluble, every state (aq)

Step 2 · Write each (aq) compound as its ions



complete ionic · no solid to keep intact

Step 3 · Cancel the spectator ions



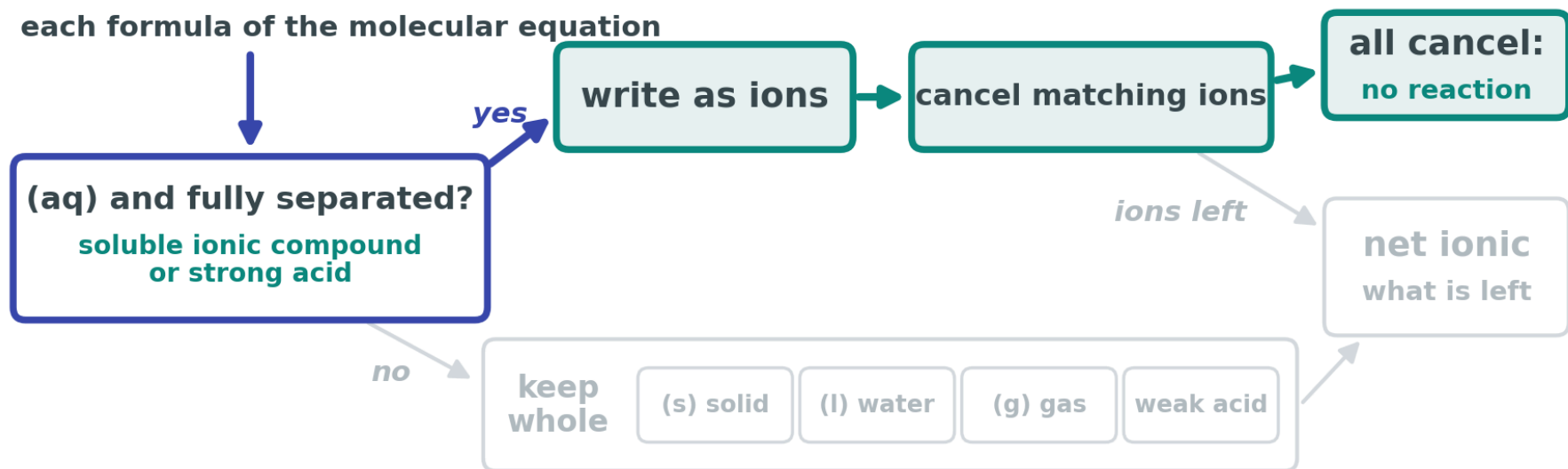
4 – 4 = 0 ions remain: no net ionic equation

 **Sense check:** Mixing produced one solution holding the same four ions. When everything cancels, no reaction occurred, which is exactly what the clear beaker showed.

✓ Worked example 3: the route on the strip



split: all four compounds · kept whole: none · found: every ion cancels



 **Sense check:** Every formula took the yes branch, so nothing was left to build a solid, water or a gas. All ions cancel: no reaction. ✓



Practice 2

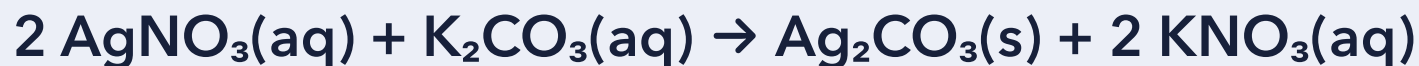


two clear solutions are mixed · a pale yellow solid forms

A few drops of silver nitrate fall into potash solution, K_2CO_3 . Which net ionic equation describes the change?

- A. $2 \text{Ag}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{Ag}_2\text{CO}_3(\text{s})$
- B. $\text{Ag}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{Ag}_2\text{CO}_3(\text{s})$
- C. $\text{K}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) \rightarrow \text{KNO}_3(\text{s})$
- D. $2 \text{Ag}^+(\text{aq}) + 2 \text{NO}_3^-(\text{aq}) + 2 \text{K}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{Ag}_2\text{CO}_3(\text{s}) + 2 \text{K}^+(\text{aq}) + 2 \text{NO}_3^-(\text{aq})$
- E. $\text{Ag}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{AgCO}_3(\text{s})$

✓ Practice 2: answer A



molecular · carbonates: insoluble except group 1 and NH_4^+ · Ag_2CO_3 : $2(1+) + 1(2-) = 0$ ✓



atoms: Ag 2 = 2 ✓ · C 1 = 1 ✓ · O 3 = 3 ✓ · charge: left $2(1+) + 1(2-) = 0$, right 0 ✓

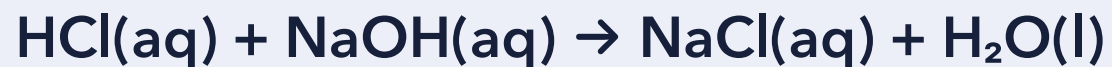
B dropped the coefficient: Ag 1 $\neq$ 2, and charge $1(1+) + 1(2-) = -1$ against 0. C precipitated the soluble pair: K^+ is group 1 and nitrate is always soluble. D is the complete ionic equation, K^+ and NO_3^- uncanceled. E copied a one-to-one ratio into the formula: AgCO_3 sums to $1(1+) + 1(2-) = -1$, not neutral.

 **Sense check:** Two 1+ silvers balance one 2– carbonate. The neutral solid sets the 2 : 1 ratio, and the charge check confirms it.



Worked example 4: acid plus base

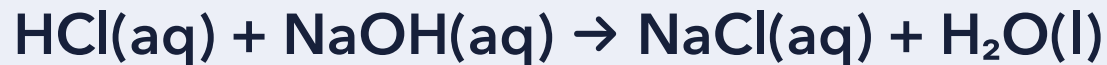
Step 1 • Write the molecular equation with states



HCl: a strong acid, all ions in water • NaOH: a soluble ionic compound • H₂O(l): a covalent liquid

Mixing hydrochloric acid with sodium hydroxide solution gives salt water: nothing visible happens, yet the beaker warms. Write the complete ionic equation, then the net ionic equation.

✓ Worked example 4: the complete ionic equation



molecular · all three (aq) compounds are strong electrolytes

Step 2 · Write each (aq) compound as its ions



charge: left $1(1+) + 1(1-) + 1(1+) + 1(1-) = 0$ · right $1(1+) + 1(1-) + 0 = 0$ ✓

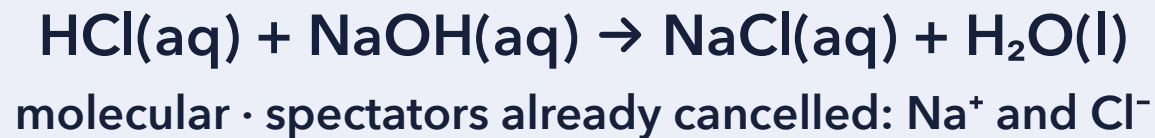
Step 3 · Cancel the spectator ions



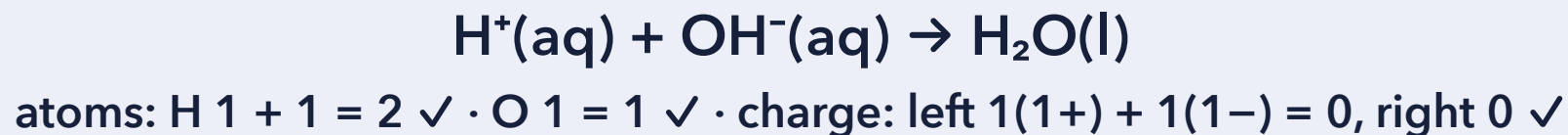
spectators: Na^+ and Cl^- · the salt of the molecular equation never left the water

 **Sense check:** No solid forms this time. The product that drives the reaction is the covalent molecule H_2O , and it keeps its H and O out of the ion pool.

✓ Worked example 4: the net ionic equation



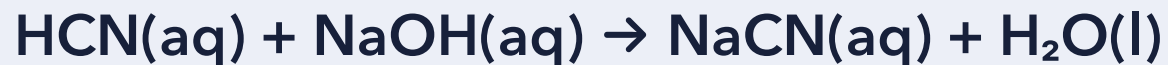
Step 4 · Check atoms and charge



Every strong acid neutralizing every soluble hydroxide gives this same net line, so long as the salt formed stays dissolved.

 **Sense check:** One net ionic equation stands behind thousands of acid-base pairs: H^+ meets OH^- and leaves the ion pool as water.

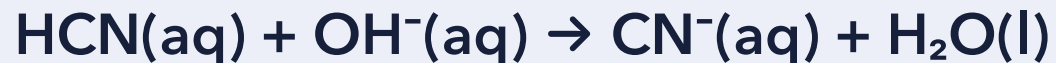
✓ Worked example 4: a weak acid instead



molecular · HCN: a weak acid, almost every molecule stays whole in water

The weak-acid case

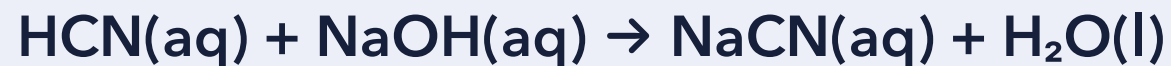
A weak acid is not written as ions: the water really holds HCN molecules.



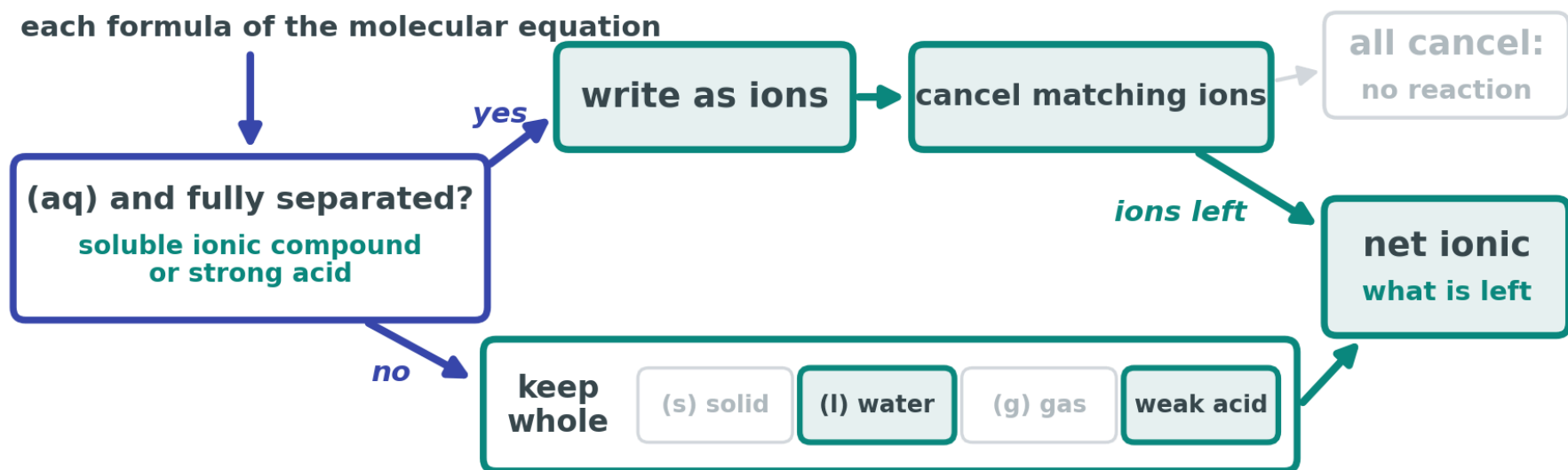
net ionic · Na⁺ is the only spectator · atoms: H 1 + 1 = 2 ✓ · C 1 = 1 ✓ · N 1 = 1 ✓ · O 1 = 1 ✓ · charge:
left 0 + 1(1–) = 1–, right 1(1–) + 0 = 1– ✓

 **Sense check:** The acid appears whole because that is what the solution contains. A net ionic equation may carry a nonzero total charge, as long as both sides match.

✓ Worked example 4: the route on the strip



split: NaOH, NaCN · kept whole: HCN (weak acid), H₂O(l) · found: $\text{HCN(aq)} + \text{OH}^{\text{-(aq)}} \rightarrow \text{CN}^{\text{-(aq)}} + \text{H}_2\text{O(l)}$

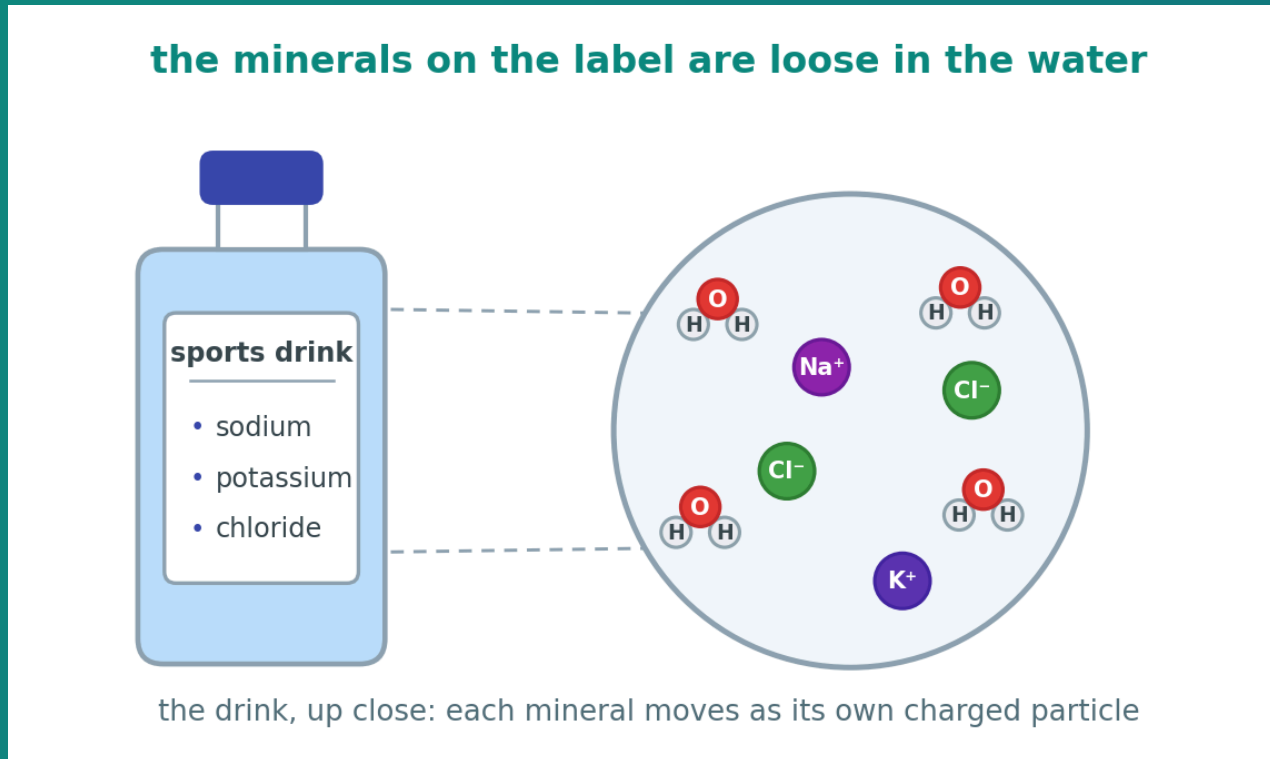


Sense check: With HCl, only water takes the no branch. With HCN, the weak acid takes it too, so the acid enters the net ionic equation whole. ✓

5 • Electrolytes & Dissociation

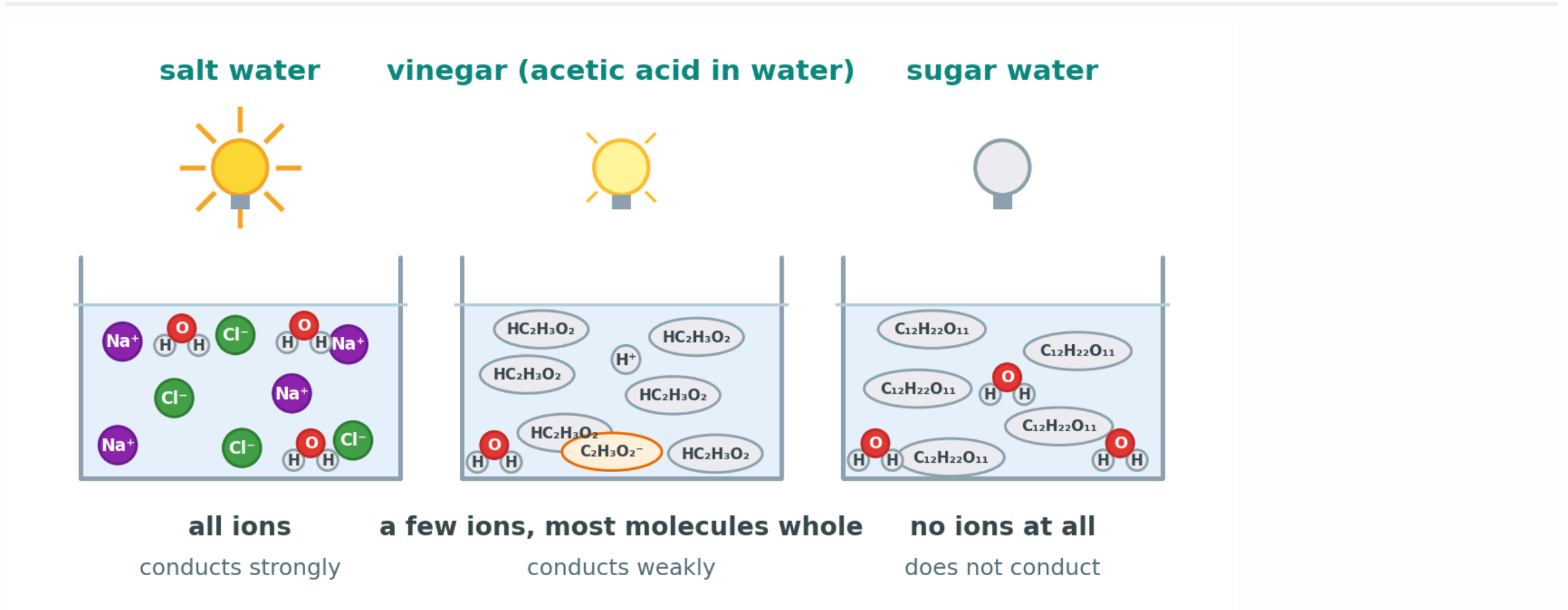
Classify a solute as a strong electrolyte, a weak electrolyte, or a nonelectrolyte from its compound type, and write its dissociation equation with the right ions, coefficients, and charge sum.

Inside a sports drink



The label lists sodium, potassium, chloride. In the bottle, each one travels through the water as a separate charged particle. Nerve and muscle signals run on these moving charges.

Conduction needs moving charges



A solution conducts only if charged particles can move through it. What a solute becomes in water sets how strongly its solution conducts: all ions, a few ions, or no ions at all.

Three classes of solute

strong electrolyte: dissolves entirely as ions

soluble ionic compounds (NaOH, KOH included) · the seven strong acids: HCl · HBr · HI · HNO₃ · H₂SO₄ · HClO₄ · HClO₃

weak electrolyte: a small fraction ionizes

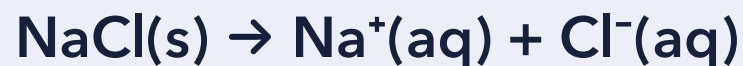
weak acids and weak bases: HC₂H₃O₂ · NH₃; most molecules stay whole

nonelectrolyte: dissolves as whole molecules

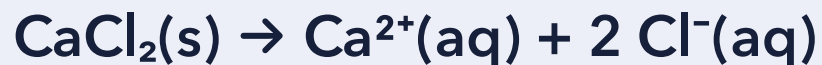
other molecular compounds: sugar · ethanol; no ions, no conduction

An **electrolyte** releases ions in water, and its solution conducts. Compound type assigns the class. Acids and bases get strength lists of their own.

Dissociation equations



1 + 1 = 2 ions per formula unit · charge: (1+) + (1−) = 0



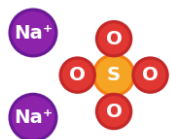
1 + 2 = 3 ions per formula unit · charge: (2+) + 2(1−) = 0

Water pulls an ionic solid apart into its separate ions: **dissociation**. Each ion keeps its identity and its charge. A subscript counts separate ions, so it becomes a coefficient.

Polyatomic ions stay in one piece

dissociation separates the ions and keeps each one whole

one Na_2SO_4 formula unit:
ions held together



water

in water: 3 separate ions,
each one whole



X never into separate atoms

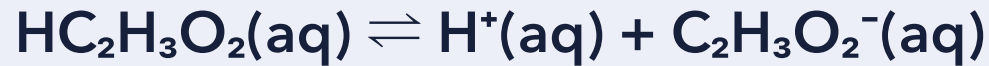


the S-O bonds do not break:
sulfate travels with its own 2- charge



Dissociation separates cations from anions. It never breaks the bonds inside a polyatomic ion: sulfate enters the water whole, carrying its 2- charge.

Weak electrolytes: partial ionization



most molecules stay whole · each ionization: $(1+) + (1-) = 0$



a weak base: only a few molecules react · charge: $(1+) + (1-) = 0$

A molecular acid or base **ionizes**: reaction with water makes new ions. A weak one barely reacts: a few molecules ionize, the rest stay whole. The double arrow marks an incomplete reaction.



The method

1. **Classify the solute.** Soluble ionic and strong acids: strong. Other acids and bases: weak. Other molecular: nonelectrolyte.
2. **Write what water makes.** Separated ions, a few ions, or whole molecules.
3. **Check the charge sum.** The ions must total zero.



Worked example 1: magnesium chloride

MgCl₂(s) dissolved in water

given: a soluble ionic compound · wanted: class + the dissociation equation

Road crews spread MgCl₂ as a de-icer, and it dissolves freely. Classify it and write the dissociation equation.

✓ Worked example 1: solution

$\text{MgCl}_2(\text{s})$ dissolved in water

given: a soluble ionic compound · wanted: class + the dissociation equation

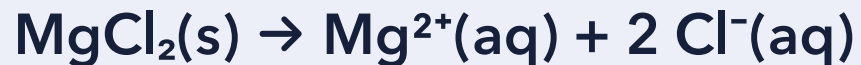
Step 1 · Classify the solute

Step 2 · Write what water makes



subscript 2 → coefficient $2 \cdot 1 + 2 = 3$ ions per formula unit

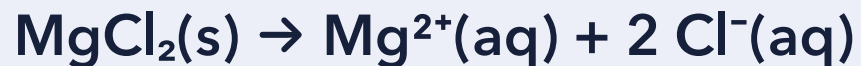
Step 3 · Check the charge sum



3 ions · charge: $(2+) + 2(1-) = 0$ ✓

 **Sense check:** The solid is neutral, so the ions it releases must cancel: one 2+ against two 1−. A nonzero sum marks a wrong formula or a wrong coefficient. ✓

✓ Worked example 1: the route on the chart



MgCl_2 : ionic, soluble · found: strong electrolyte, 3 ions per formula unit



🧠 **Sense check:** Two yes answers: ionic, then soluble. The acid questions never come up for an ionic solute. ✓

 Worked example 2: Na_2CO_3 , NH_3 , $\text{C}_2\text{H}_5\text{OH}$ 

washing soda · household ammonia · ethanol · wanted: each class + what each solution contains

All three dissolve freely in water. Classify each and write what its solution contains.

✓ Worked example 2: classifying

$\text{Na}_2\text{CO}_3 \cdot \text{NH}_3 \cdot \text{C}_2\text{H}_5\text{OH}$
washing soda · household ammonia · ethanol

Step 1 · Classify the solute

solute	type	class
Na_2CO_3	soluble ionic compound	strong electrolyte
NH_3	molecular base, not an ionic hydroxide	weak electrolyte
$\text{C}_2\text{H}_5\text{OH}$	molecular, neither acid nor base	nonelectrolyte

 **Sense check:** All three bottles look identical. The compound type, not the appearance, separates them.

✓ Worked example 2: what each solution contains



classified: strong electrolyte · weak electrolyte · nonelectrolyte

Step 2 · Write what water makes

Step 3 · Check the charge sum



2 + 1 = 3 ions · charge: $2(1+) + (2-) = 0$ · carbonate stays one piece



a few ions · charge: $(1+) + (1-) = 0$ · most NH_3 molecules stay whole

$\text{C}_2\text{H}_5\text{OH}(\text{aq})$: dissolves as whole molecules

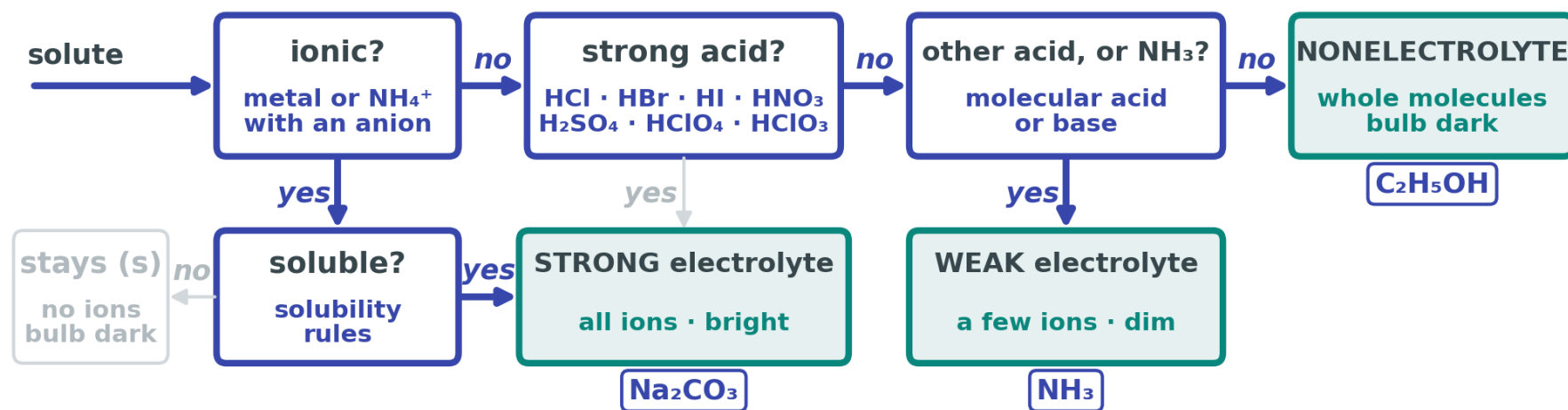
0 ions · the OH is covalently bonded, not OH^-

 **Sense check:** Three clear solutions, three bulb readings: bright, dim, dark. ✓

✓ Worked example 2: the route on the chart



Na_2CO_3 : ionic, soluble · NH_3 : not ionic, not on the acid list, a base · $\text{C}_2\text{H}_5\text{OH}$: every answer no



🧠 **Sense check:** Three solutes leave by three exits. Only the ionic one needed the solubility rules. ✓

Your turn: magnesium nitrate

$\text{Mg}(\text{NO}_3)_2(\text{s})$ dissolved in water
 given: a soluble ionic compound · nitrate: NO_3^-

step	work	result
1 · classify the solute	soluble ionic compound	_____ electrolyte
2 · write what water makes	$\text{Mg}(\text{NO}_3)_2(\text{s}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \underline{\hspace{1cm}}$	ions per formula unit: _____
3 · check the charge sum	$(2+) + 2(1-)$	= _____

Complete the classification and the equation.

$\text{Mg}(\text{NO}_3)_2(\text{s}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2 \text{NO}_3^-(\text{aq})$
 strong electrolyte · $1 + 2 = 3$ ions · charge: $(2+) + 2(1-) = 0$ · each nitrate leaves whole

⚠ Where this goes wrong

Reading a subscript as a bonded pair. CaCl_2 never releases a Cl_2^{2-} unit. The subscript counts separate ions: $\text{Ca}^{2+} + 2 \text{Cl}^-$ makes $1 + 2 = 3$ ions, not $1 + 1 = 2$.

Breaking a polyatomic ion into atoms. Na_2CO_3 gives $2 \text{Na}^+ + \text{CO}_3^{2-} = 3$ ions, never $2 + 1 + 3 = 6$ pieces. Dissociation separates ions; it does not break the bonds inside one.

Calling sugar a weak electrolyte. Weak means a few ions form. Sugar forms none: its solution conducts no better than pure water. Nonelectrolyte.

Reading a molecular OH as hydroxide. Ethanol's OH is covalently bonded and stays put. Only ionic hydroxides such as NaOH release OH^- .



Practice 1

K_3PO_4 dissolved in water

given: a soluble ionic compound · phosphate: PO_4^{3-}

Fertilizer-grade potassium phosphate dissolves freely in water. Which statement classifies it and describes what its solution contains?

- A. Weak electrolyte: a salt built around a polyatomic ion dissociates only partially
- B. Strong electrolyte: it dissociates completely into 3K^+ and PO_4^{3-} , four ions per formula unit
- C. Strong electrolyte: it dissociates completely into K_3^+ and PO_4^{3-} , two ions per formula unit
- D. Nonelectrolyte: it dissolves as intact, neutral K_3PO_4 molecules

✓ Practice 1: answer B



$$3 + 1 = 4 \text{ ions} \cdot \text{charge: } 3(1+) + (3-) = 0$$

A: solubility decides, not the anion; a soluble salt dissociates completely, polyatomic ion or not. C: the subscript counts three separate K^+ ions; no K_3^+ unit exists, and $1 + 1 = 2$ undercounts the ions. D: an ionic compound has no molecules; only separated ions enter the water.

 **Sense check:** Four ions from one formula unit, and the charges cancel: $3(1+) + (3-) = 0$. ✓



Worked example 3: ion concentrations

0.100 M AlBr_3 dissolved in water

M, molarity = mol of solute per liter of solution · wanted: the molarity of Br^-

Aluminum bromide dissolves freely. Every formula unit that dissolves releases its ions into the same liter. Find the concentration of Br^- in the solution.

✓ Worked example 3: solution

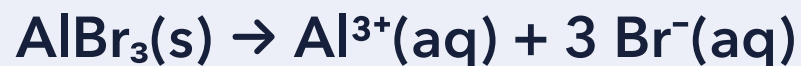


M = mol per liter of solution

Step 1 · Classify the solute

Step 2 · Write what water makes

Step 3 · Check the charge sum



1 + 3 = 4 ions · charge: (3+) + 3(1-) = 0 ✓

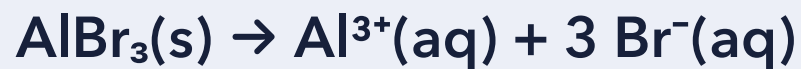
Ion molarity · multiply by the coefficient

Each liter holds 0.100 mol of dissolved AlBr_3 , and each mole releases 3 mol of Br^- .

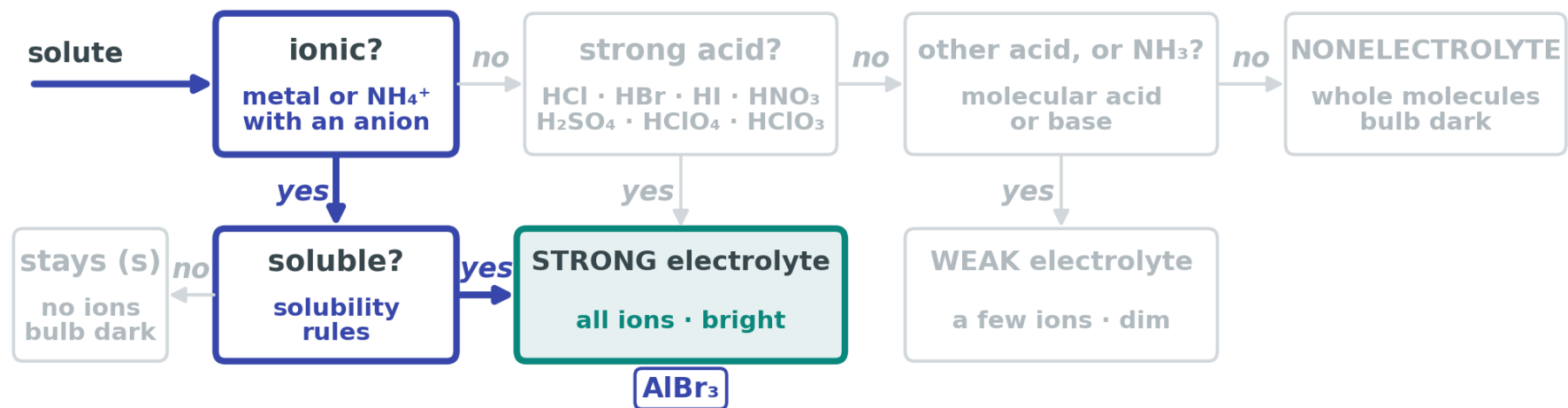
$$0.100 \text{ mol AlBr}_3 \times \frac{3 \text{ mol Br}^-}{1 \text{ mol AlBr}_3} = \mathbf{0.300 \text{ mol Br}^-} \text{ per liter} = 0.300 \text{ M}$$

 **Sense check:** The subscript became a concentration ratio: 0.300 M Br^- against 0.100 M Al^{3+} .

✓ Worked example 3: the route on the chart



AlBr_3 : ionic, soluble · found: strong electrolyte · $0.100 \text{ M AlBr}_3 \rightarrow 0.300 \text{ M Br}^{-}$



Sense check: The strong exit means every formula unit dissociates. Only then does the coefficient 3 turn 0.100 M AlBr_3 into 0.300 M Br^{-} . ✓



Practice 2

$\text{Fe}_2(\text{SO}_4)_3$ dissolved in water

given: 0.60 M SO_4^{2-} required · wanted: the molarity of $\text{Fe}_2(\text{SO}_4)_3$

A water plant's dosing tank must reach 0.60 M sulfate ion, supplied by dissolving iron(III) sulfate. What molarity of $\text{Fe}_2(\text{SO}_4)_3$ does the tank need?

- A. 0.30
- B. 1.8
- C. 0.60
- D. 0.20

✓ Practice 2: answer D



$$2 + 3 = 5 \text{ ions} \cdot \text{charge: } 2(3+) + 3(2-) = 0 \checkmark$$

$$0.60 \text{ mol SO}_4^{2-} \times \frac{1 \text{ mol Fe}_2(\text{SO}_4)_3}{3 \text{ mol SO}_4^{2-}} = \mathbf{0.20 \text{ mol Fe}_2(\text{SO}_4)_3 \text{ per liter} = 0.20 \text{ M (answer D)}}$$

C skipped the ratio: $0.60 \times 1 = 0.60 \text{ M}$ assumes one sulfate per formula unit. B flipped the ratio: $0.60 \times 3 = 1.8 \text{ M}$, a tank at $1.8 \times 3 = 5.4 \text{ M}$ sulfate. A used iron's subscript: $0.60 \div 2 = 0.30 \text{ M}$.

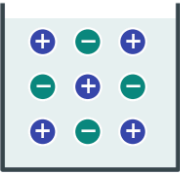
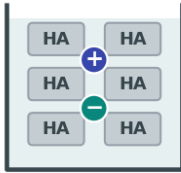
 **Sense check:** Each formula unit releases three sulfates, so the salt runs at one third of the target: $3 \times 0.20 = 0.60 \text{ M}$ sulfate ✓.

6 · Acids and Bases, Strong vs Weak

Classify any acid or base as strong or weak using the memorize-lists, write its ionization with the correct arrow ($\rightarrow$ for full, $\rightleftharpoons$ for partial), and predict the products of neutralization and gas-evolution reactions.

Strong is not the same as a lot

Strong vs Weak: the degree of ionization, not the amount

STRONG ionizes ~100% → strong electrolyte	WEAK ionizes partially ⇌ weak electrolyte
Strong acids: memorize these 7 HCl · HBr · HI (hydrohalic) HNO ₃ · H ₂ SO ₄ · HClO ₄ · HClO ₃ (oxoacids)	Weak acids: everything else CH ₃ COOH · HF · H ₂ CO ₃ · H ₃ PO ₄ (acetic, hydrofluoric, carbonic...) HCN · HNO ₂ · HClO · H ₂ S (HClO weak ⇌ HClO ₃ , HClO ₄ strong)
Strong bases: 5 + 3 = 8 group 1: LiOH NaOH KOH RbOH CsOH group 2 (heavy): Ca(OH) ₂ Sr(OH) ₂ Ba(OH) ₂	Weak bases NH ₃ (ammonia) CH ₃ NH ₂ and other amines
HCl → H⁺ + Cl⁻  every molecule splits into ions (≈ 100% ionized)	CH₃COOH ⇌ H⁺ + CH₃COO⁻  mostly intact molecules, few ions
Concentrated ≠ strong. Strength = the fraction that ionizes; concentration = how much is dissolved.	

Stomach acid is hydrochloric acid, HCl, a *strong* acid, yet dilute. Strong and weak describe how completely an acid ionizes, not how much of it is dissolved.

What makes something an acid or a base

Dissolved in water, an **acid** produces H^+ , carried as the hydronium ion H_3O^+ . A **base** produces OH^- .

acid $\rightarrow \text{H}^+$ (H_3O^+) in water · base $\rightarrow \text{OH}^-$ in water

Arrhenius definitions · Brønsted-Lowry widens them: an acid donates a proton (H^+), a base accepts one

Every acid here has an H to release; every hydroxide carries its OH^- .

Strong vs weak = the degree of ionization

Strong means essentially every molecule ionizes; **weak** means a small fraction. The $\rightleftharpoons$ arrow marks the reverse reaction: ions rejoin as fast as molecules split, so the mixture settles mostly intact.

$\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$ ($\approx 100\%$, one $\rightarrow$ arrow) · $\text{CH}_3\text{COOH} \rightleftharpoons \text{H}^+ + \text{CH}_3\text{COO}^-$ (partial, $\rightleftharpoons$)
strong = fully ionized = a strong electrolyte · weak = barely ionized = a weak electrolyte

What a 1 M acid bottle contains

a bottle labeled 1 M HCl actually has no HCl in it, but 1 M $\text{H}^+(\text{aq})$ and 1 M $\text{Cl}^-(\text{aq})$
strong acid: every molecule ionized · nothing left to match the label

a bottle labeled 1 M HA, where HA is a weak acid, actually has mostly HA in it and
only a small amount of $\text{H}^+(\text{aq})$ and $\text{A}^-(\text{aq})$
weak acid: the molecules survive · few ions

The label names what was dissolved. The strong or weak classification tells what the water holds now.

The seven strong acids: memorize them

Only **seven** common acids are strong; everything not on the list is weak. The list is measured, not deduced: these seven ionize essentially completely in water.



3 hydrohalic (HCl, HBr, HI) + 4 oxoacids (HNO₃, H₂SO₄, HClO₄, HClO₃) = 3 + 4 = 7

not on it: CH₃COOH · HF · H₂CO₃ · H₃PO₄ → all weak, partial ionization ($\rightleftharpoons$)
the structural reason for the split waits for equilibrium

Strong bases, and neutralization

The strong bases are the **metal hydroxides that ionize completely**: every group 1 hydroxide, plus the heavy group 2 hydroxides.

group 1: LiOH NaOH KOH RbOH CsOH · group 2 (heavy): Ca(OH)₂ Sr(OH)₂ Ba(OH)₂
5 + 3 = 8 strong bases · Ca(OH)₂ → Ca²⁺ + 2 OH⁻ releases 2 × 1 = 2 hydroxide ions · Ca(OH)₂ and Sr(OH)₂ barely dissolve (the chart says insoluble), but all that dissolves ionizes fully



ammonia and amines make OH⁻ only partially: weak bases · an acid and a base neutralize to a salt plus water

Ammonia and the amines are **weak** bases. An acid and a base **neutralize** each other.



The method

1. **Acid or base?** An H to release: acid. Hydroxide or NH_3 /amine: base.
2. **On a memorize list?** Listed: strong. Not listed: weak.
3. **Pick the arrow.** Strong ionizes fully: $\rightarrow$. Weak, partially: $\rightleftharpoons$.
4. **Ignore concentration.** Strength is the fraction ionized.



Worked example 1: classify three species



for each: strong or weak? which electrolyte? which arrow: $\rightarrow$ or $\rightleftharpoons$?

Nitric acid, hydrofluoric acid, and calcium hydroxide. Check each against the memorize-lists, then write how it behaves in water.

✓ Worked example 1: solution

HNO₃ · on the strong-acid list



full ionization, single $\rightarrow$ arrow · $1 + 1 = 2$ ions per formula unit

HF · not on the list $\rightarrow$ weak



mostly intact HF molecules · the $\rightleftharpoons$ arrow marks partial ionization

Ca(OH)₂ · a heavy group 2 hydroxide $\rightarrow$ strong base



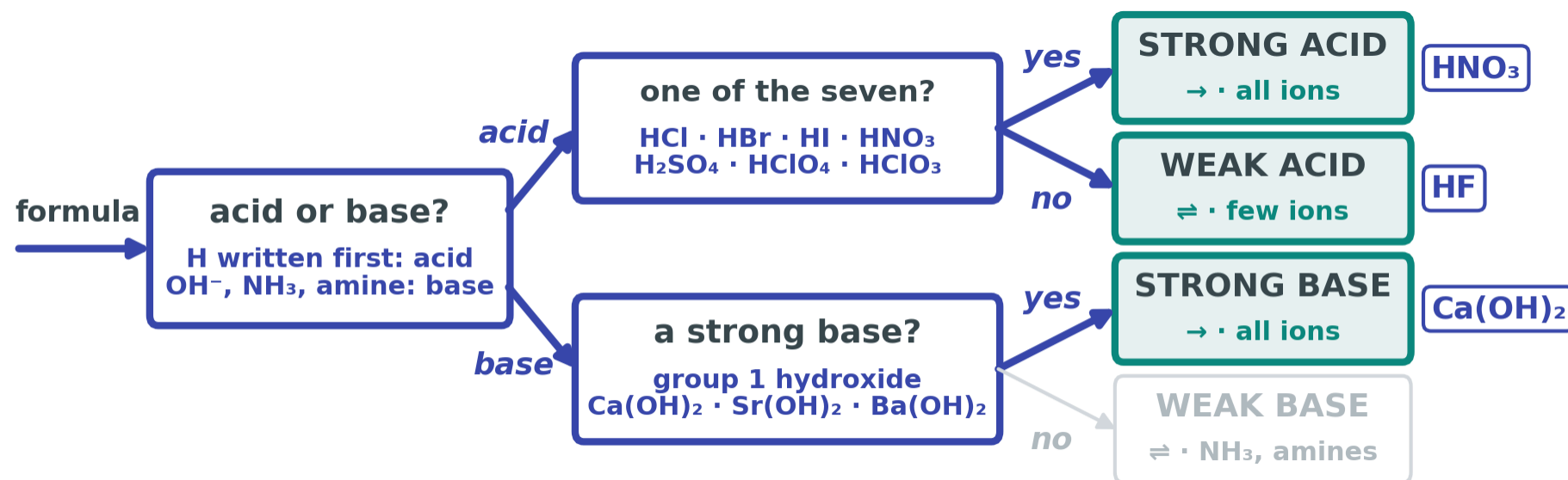
strong base, strong electrolyte · $1 + 2 = 3$ ions, of which $2 \times 1 = 2$ are OH⁻

 **Sense check:** On the lists, HNO₃ and Ca(OH)₂ ionize fully ($\rightarrow$); HF is off the list, mostly intact ($\rightleftharpoons$).

✓ Worked example 1: the route on the chart



HNO_3 : acid, one of the seven · HF : acid, not one of the seven · Ca(OH)_2 : base, a strong base



🧠 **Sense check:** Three formulas leave by three exits. Only the two lists decided strong or weak. ✓

⚠ Where this goes wrong

"HF is strong because F is so electronegative." Electronegativity does not decide it: HF is simply not on the list of seven, so it is weak ($\rightleftharpoons$).

"Concentrated means strong." Concentration is amount dissolved; strength is the fraction ionized. 10 M acetic acid is still weak; 0.01 M HCl is still strong.

Sulfuric $\leftrightarrow$ acetic mix-up. H_2SO_4 is on the list (strong, $\rightarrow$); acetic acid CH_3COOH is not (weak, $\rightleftharpoons$). Being an acid does not make it strong.

" NH_3 can't be a base: it has no OH." It makes OH^- by pulling H^+ off water: $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$: a weak base.

Practice 1: spot the exception

Three of these ionize essentially 100% in water. Which one is **not** a strong acid?

- A. HClO_4
- B. HCl
- C. HClO
- D. HNO_3

✓ Practice 1: answer C

$\text{HClO} \rightleftharpoons \text{H}^+ + \text{ClO}^-$: a weak acid (answer C)

HClO (hypochlorous) is not among the seven; HClO_4 , HCl , HNO_3 all are

Watch the chlorine oxoacids: HClO_4 (perchloric) and HClO_3 (chloric) are on the strong list, but HClO (hypochlorous) is **weak**: it ionizes only partially ($\rightleftharpoons$). HCl and HNO_3 are strong ($\rightarrow$).

 **Sense check:** Two fewer oxygens separate HClO_3 , strong, from HClO , weak. Classify from the list, not from the look of the formula.



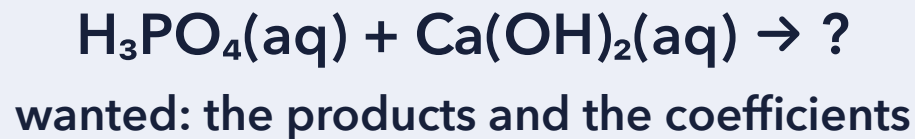
Worked example 2: predict and balance a neutralization



acid + base $\rightarrow$ salt + water · wanted: the products and the coefficients

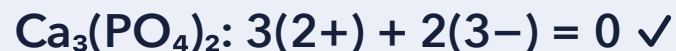
Phosphoric acid meets calcium hydroxide. One HCl releases one H^+ : monoprotic. H_2SO_4 releases two: diprotic. H_3PO_4 releases three: triprotic. Predict both products, then balance.

✓ Worked example 2: the products



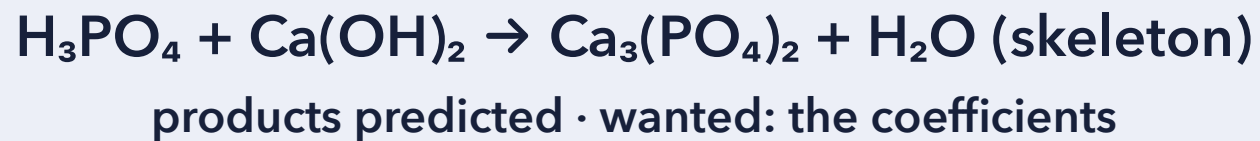
Predict the products

The salt pairs Ca^{2+} with PO_4^{3-} ; charge balance builds $\text{Ca}_3(\text{PO}_4)_2$. The other product of every neutralization is water.

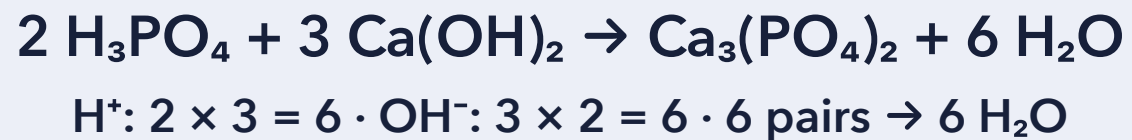


 **Sense check:** The salt's formula comes from the ion charges, never from the coefficients: the products must be neutral before any balancing starts.

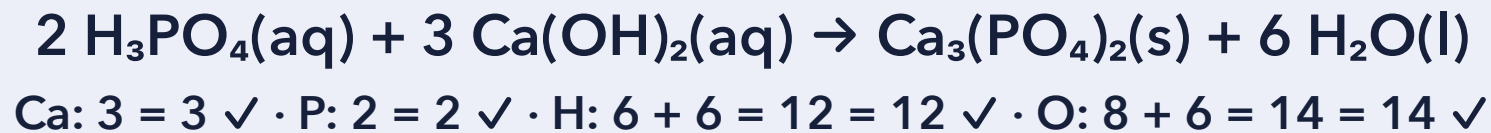
✓ Worked example 2: the coefficients



Match H^+ to OH^-



Recheck every count

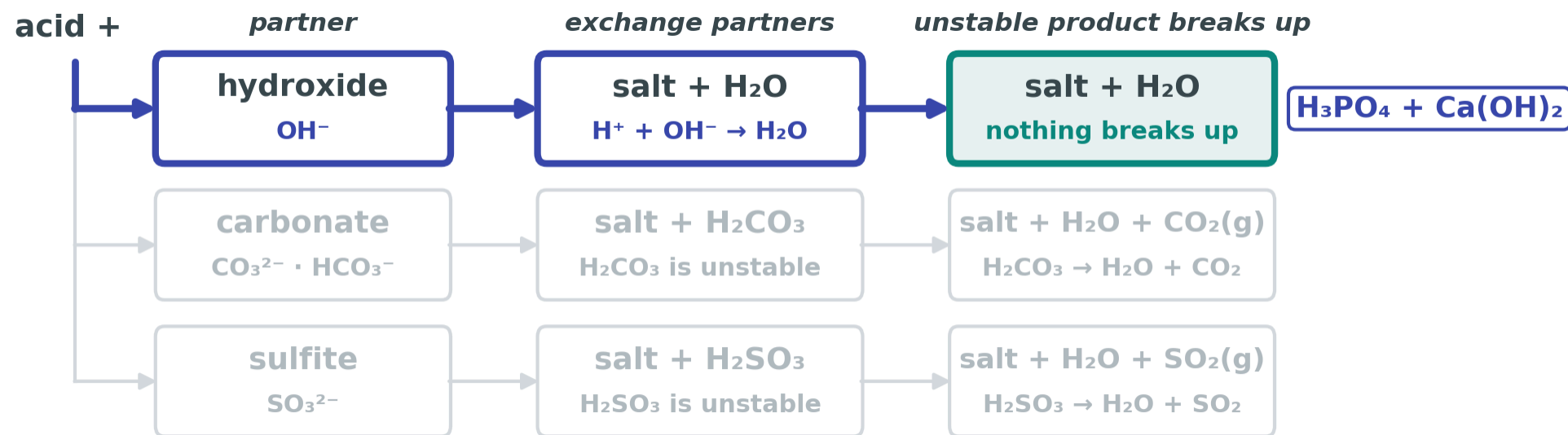


 **Sense check:** The coefficients 2 and 3 came from matching six H^+ to six OH^- , and one water forms per pair: $6 \text{H}_2\text{O}$.

✓ Worked example 2: the route on the chart



partner: a hydroxide · exchange: salt + water · nothing breaks up



🧠 **Sense check:** A hydroxide partner stops after the exchange: a salt and water, no gas. ✓



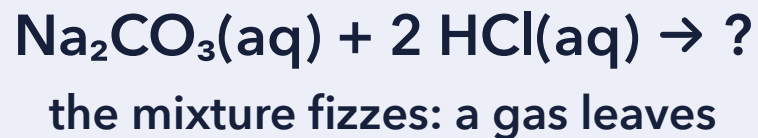
Worked example 3: acid plus a carbonate



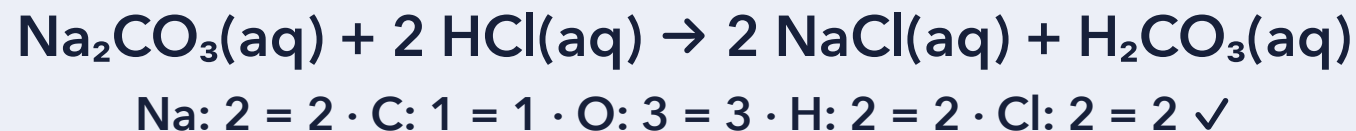
wanted: the products, found in two steps

Washing soda meets hydrochloric acid, and the mixture fizzes. The gas is the product to explain. Work the exchange first, then watch what the exchange product does.

✓ Worked example 3: the exchange



First reaction • exchange partners



The exchange is ordinary. The product H_2CO_3 , carbonic acid, is not: it cannot survive in water.

 **Sense check:** On paper the swap looks routine. The fizz says otherwise: one of these products refuses to stay in the water.

✓ Worked example 3: the gas appears



the exchange product H_2CO_3 cannot survive in water

Second reaction • the unstable product breaks up



H: 2 = 2 • C: 1 = 1 • O: 3 = 1 + 2 = 3 ✓ • the fizz is CO_2 leaving

Overall



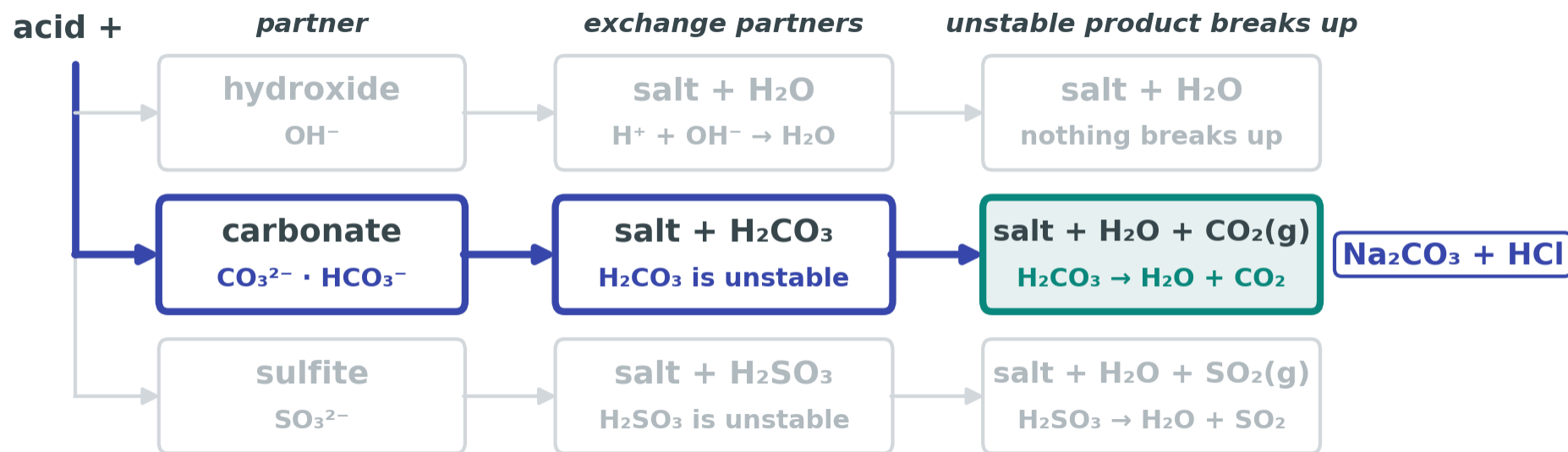
Na: 2 = 2 • C: 1 = 1 • O: 3 = 3 • H: 2 = 2 • Cl: 2 = 2 ✓

 **Sense check:** Acid plus a carbonate or bicarbonate always ends here: CO_2 , water, and a salt. The two-step explains the fizz that one overall line hides.

✓ Worked example 3: the route on the chart



partner: a carbonate · exchange: $\text{NaCl} + \text{H}_2\text{CO}_3$ · H_2CO_3 breaks up



🧠 **Sense check:** A carbonate partner runs the full lane: the exchange, then the breakup into water and CO_2 . ✓

Three products that never survive



memorize all three · each breaks into water plus a gas the moment it forms

Three exchange products are unstable. When one appears among the products, replace it with its water and gas. Acid plus a carbonate or bicarbonate always gives CO_2 , water, and a salt.

Your turn: acid plus a sulfite



exchange first · then check the unstable-products list

step	your work
exchange partners	2 KCl + _____
unstable product breaks up	$\text{H}_2\text{SO}_3 \rightarrow \text{H}_2\text{O}(\text{l}) + \text{_____}$
overall equation	$\text{K}_2\text{SO}_3(\text{aq}) + 2 \text{HCl}(\text{aq}) \rightarrow 2 \text{KCl}(\text{aq}) + \text{_____} + \text{_____}$

Complete all three lines.



 **Sense check:** An acid meeting a sulfite releases SO_2 , the sharp smell of a struck match.



Practice 4



baking soda + vinegar's acetic acid

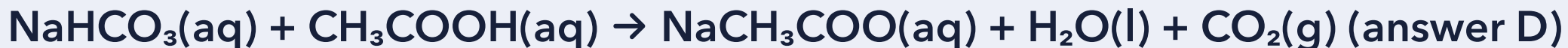
Baking soda is stirred into vinegar. Which products form?

- A. $\text{NaCH}_3\text{COO}(\text{aq}) + \text{H}_2\text{CO}_3(\text{aq})$
- B. $\text{NaCH}_3\text{COO}(\text{aq}) + \text{H}_2(\text{g})$
- C. No reaction: acetic acid is weak, so its molecules stay whole
- D. $\text{NaCH}_3\text{COO}(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

✓ Practice 4: answer D



first reaction: exchange partners · H_2CO_3 is on the never-survive list



Na: 1 = 1 · C: 1 + 2 = 3 = 2 + 1 · H: 1 + 4 = 5 = 3 + 2 · O: 3 + 2 = 5 = 2 + 1 + 2 ✓

A stopped at the exchange: H_2CO_3 breaks into H_2O and CO_2 the moment it forms. B expects H_2 , the gas an acid releases from an active metal; a bicarbonate trades partners instead. C reads weak as unreactive: weak describes the fraction ionized, and acid plus a carbonate or bicarbonate always gives CO_2 , water, and a salt.

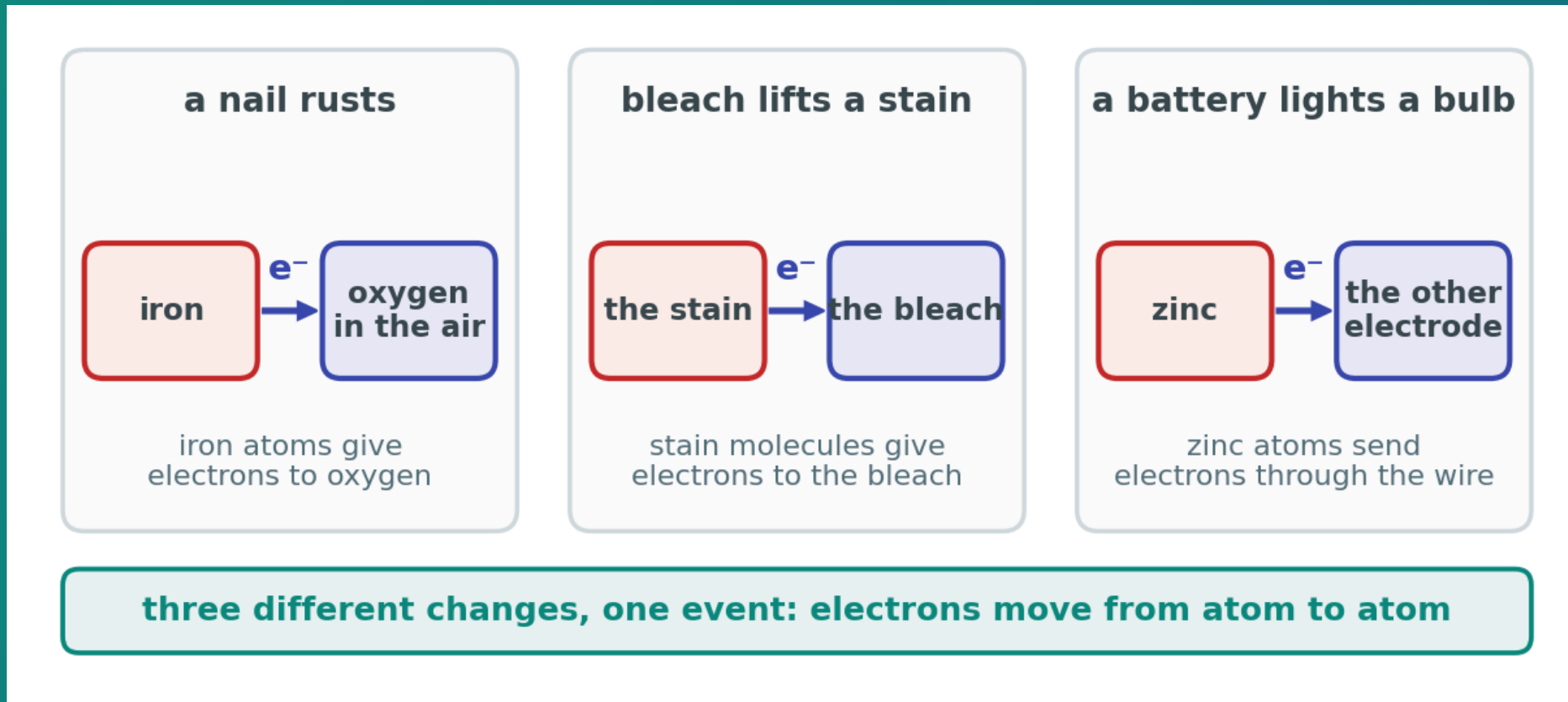


Sense check: The fizz in the cup is the CO_2 from the H_2CO_3 breakup, even with a weak acid.

7 • Oxidation Numbers

Assign oxidation numbers by the priority rules, and find any element the rules skip by setting the sum equal to the species' charge.

Rust, bleach, and batteries



A nail rusts. Bleach lifts a stain. A battery lights a bulb. In each change, atoms give electrons to other atoms. A number on each atom keeps count.

Charges in a formula sum to zero

$\text{Fe}_2\text{O}_3 \rightarrow$ iron(III) oxide

three O^{2-} : $3(-2) = -6$ · split over two Fe: $6 \div 2 = 3+$ each · $2(+3) + 3(-2) = 0 \checkmark$

Naming rust used this sum: the oxide charges set iron's charge. Oxidation numbers run the same sum on every atom, even in molecules such as CO_2 , where no atom is a real ion.

The oxidation number



$$2(+1) + 1(-2) = 0 \checkmark \cdot \text{the numbers sum to the compound's charge}$$

An atom's **oxidation number** is the charge it would carry if the shared electrons were assigned by rule. The rules form a priority list. One check governs every assignment: the numbers sum to the species' charge.

The rules, in priority order

work down: the first rule that applies wins

1 an element by itself **0** Fe · O₂ · H₂

2 a monatomic ion **its charge** Ca²⁺ → +2

3 fluorine **-1** in every compound

4 Group 1 · Group 2 **+1 · +2** Na → +1 · Mg → +2

5 hydrogen **+1** -1 with a metal: NaH

6 oxygen **-2** -1 in a peroxide: H₂O₂

then: the numbers sum to the species' charge (neutral 0 · ion its charge)

The list is one idea: assign shared electrons to the more electronegative atom. O nearly always wins, hence -2; a metal never beats H, so NaH's H is -1.

Oxidation number and ionic charge



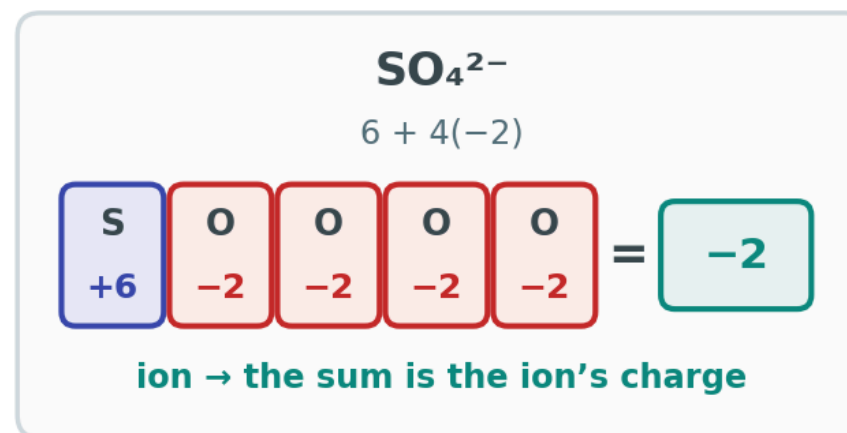
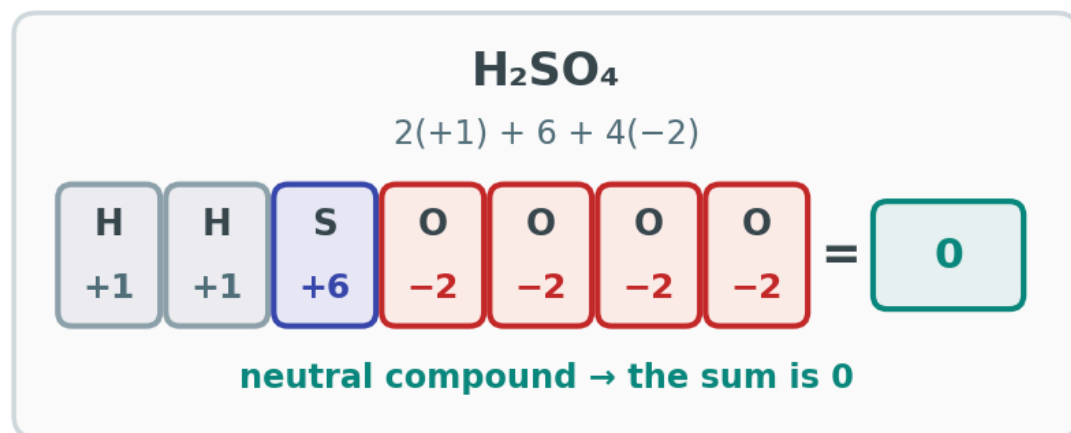
a monatomic ion: the number is the ion's real charge



$4 + 2(-2) = 0 \checkmark$ · no C^{4+} ion here; the atoms share electrons

A monatomic ion's oxidation number is its charge. In a molecule, no atom holds a full charge; the number records assigned electrons. It is written sign-first: +4, not 4+.

What the sum must equal



S is +6 in both: the sum's target changes, not the rules

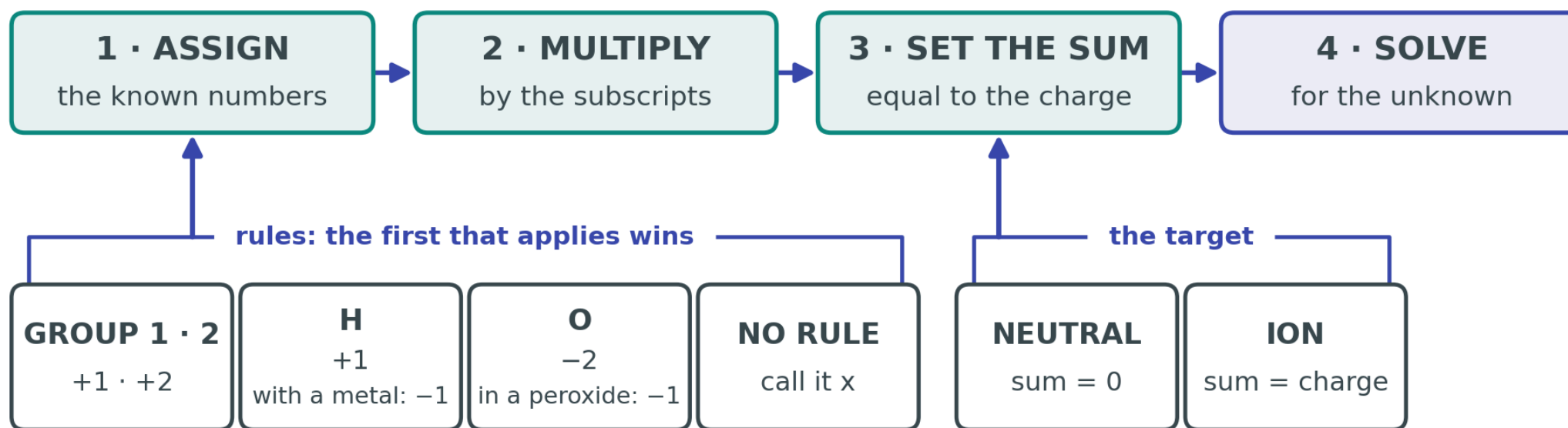
A neutral compound's numbers sum to zero. A polyatomic ion's numbers sum to the charge written on the ion, never to zero. The target changes; the rules never do.



The method

1. **Assign the known numbers.** The first rule that applies wins.
2. **Multiply by the subscripts.**
3. **Set the sum equal to the charge.** Neutral compound: 0. Polyatomic ion: its charge.
4. **Solve for the unknown.**

One route for every oxidation number



Every problem runs the same four steps. The rule cards supply the known numbers; the species' charge sets the target. A solution lights only the cards it used.

 Guided example: CH₄

given: a neutral compound · wanted: every oxidation number

Natural gas is mostly methane. Assign an oxidation number to each element.

✓ Guided example: solution



given: a neutral compound · wanted: every oxidation number

Step 1 · Assign the known numbers

Step 2 · Multiply by the subscripts

Step 3 · Set the sum equal to the charge

$$x + 4(+1) = 0$$

C: x · H: $+1$ each, 4 atoms $\rightarrow +4$ · neutral $\rightarrow$ sum 0

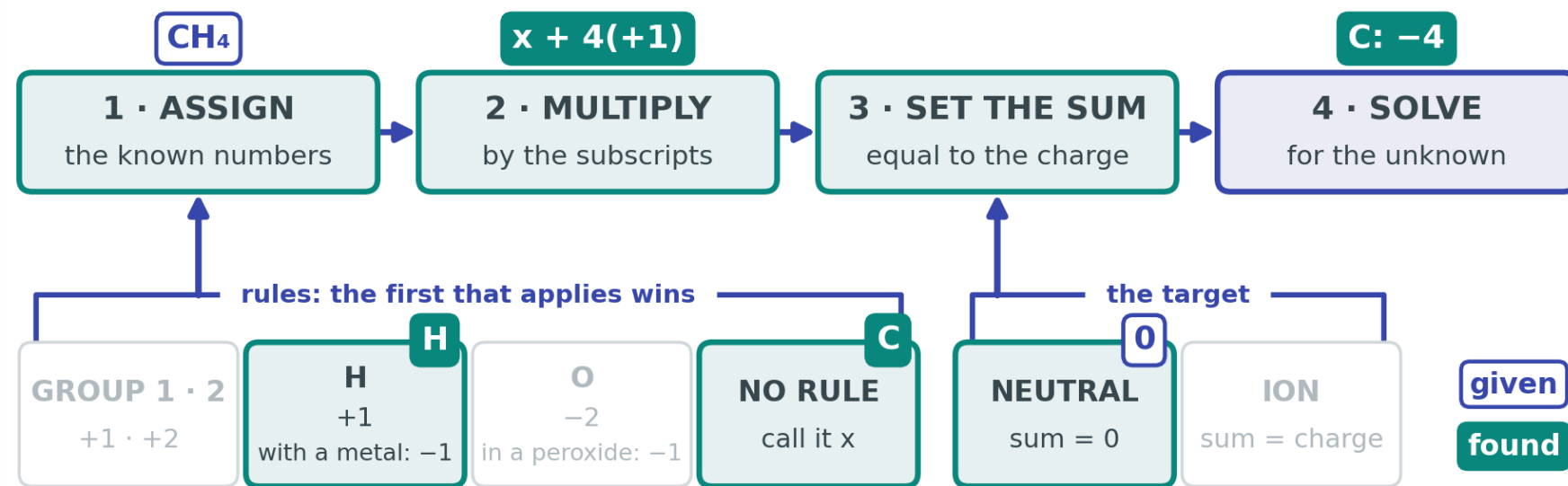
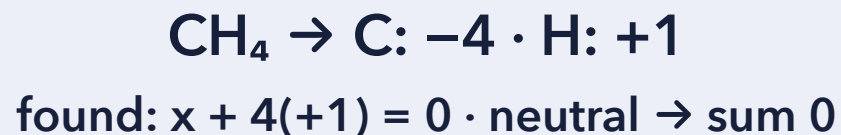
Step 4 · Solve for the unknown

$$x = 0 - 4 = -4$$

C: -4 · H: $+1 \times 4$ · check: $-4 + 4(+1) = 0$ ✓

 **Sense check:** H holds the positive side, so C balances from the negative side. The unknown can land negative.

✓ Guided example: the route on the map



🧠 **Sense check:** Only the H card applied. C had no rule, so the sum to 0 set its number. ✓



Worked example 1: SO_2



given: a neutral compound · wanted: every oxidation number

Sulfur dioxide forms when coal burns. Assign an oxidation number to each element.

✓ Worked example 1: solution



given: a neutral compound · wanted: every oxidation number

Step 1 · Assign the known numbers

Step 2 · Multiply by the subscripts

Step 3 · Set the sum equal to the charge

$$x + 2(-2) = 0$$

S: x · O: -2 each, 2 atoms $\rightarrow -4$ · neutral $\rightarrow$ sum 0

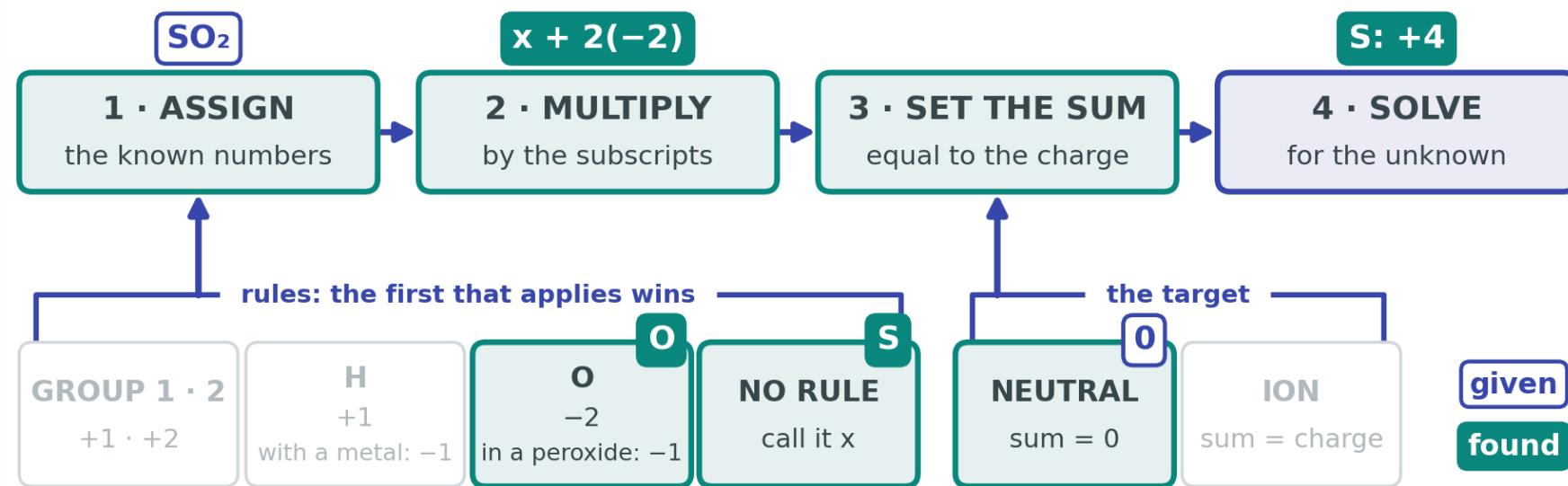
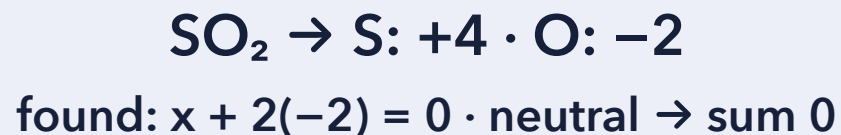
Step 4 · Solve for the unknown

$$x = 0 + 4 = +4$$

S: $+4$ · O: $-2, -2$ · check: $+4 + 2(-2) = 0$ ✓

 **Sense check:** The numbers sum to 0, a neutral compound's charge. No S^{4+} ion exists in SO_2 ; $+4$ records assigned electrons, not a real charge.

✓ Worked example 1: the route on the map



🧠 **Sense check:** The O card gave the known numbers. S took what the sum to 0 required. ✓



Worked example 2: KMnO_4



given: a neutral compound · wanted: the oxidation number of Mn

Potassium permanganate is a deep purple disinfectant. Two of its three elements have rules. Find the oxidation number of Mn.

✓ Worked example 2: solution



given: a neutral compound · wanted: the oxidation number of Mn

Step 1 · Assign the known numbers

Step 2 · Multiply by the subscripts

Step 3 · Set the sum equal to the charge

$$(+1) + x + 4(-2) = 0$$

K: +1 · Mn: x · O: -2 each, 4 atoms → -8 · neutral → sum 0

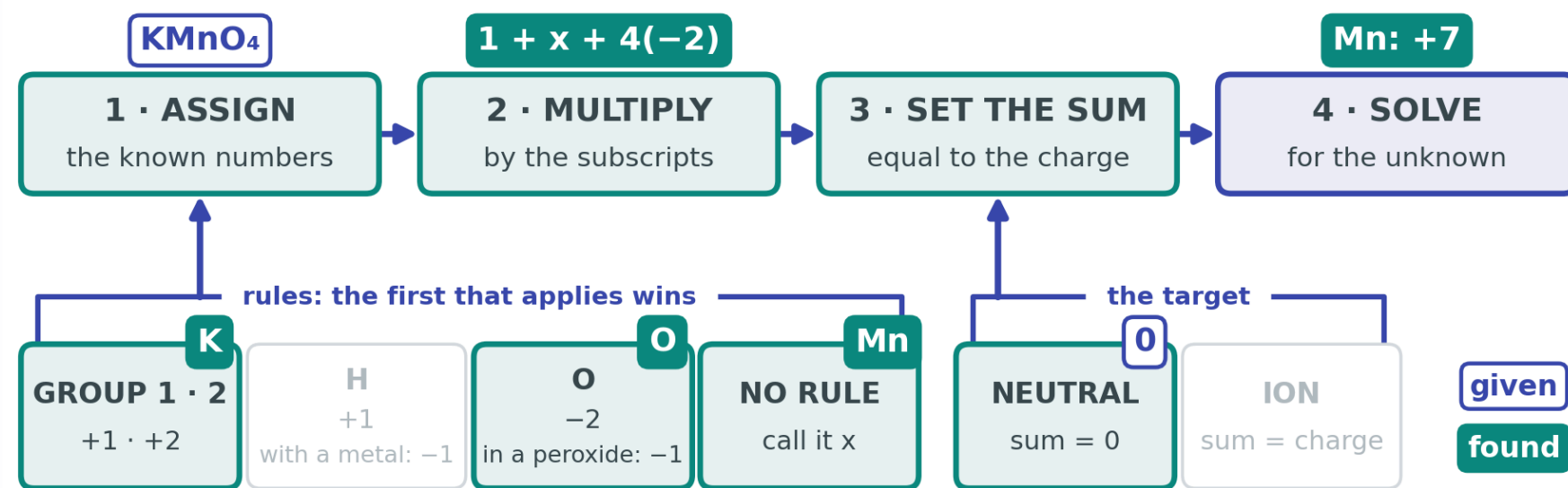
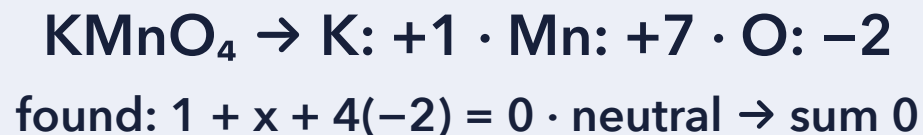
Step 4 · Solve for the unknown

$$x = 0 - 1 + 8 = +7$$

K: +1 · Mn: +7 · O: -2 × 4 · check: +1 + 7 + 4(-2) = 0 ✓

🧠 **Sense check:** +1 + 7 - 8 = 0 ✓. An element without a rule of its own gets its number from the sum.

✓ Worked example 2: the route on the map



🧠 **Sense check:** Two rule cards applied: Group 1 for K, and O. Mn took what the sum to 0 required. ✓

Your turn: H_2SO_4



given: a neutral compound · wanted: the oxidation number of S

step	work
1 · assign the known numbers	H: +1 · O: -2 · S: x
2 · multiply by the subscripts	$2(+1) = +2 \cdot 4(-2) = \underline{\hspace{2cm}}$
3 · set the sum equal to the charge	$+2 + x + (-8) = \underline{\hspace{2cm}}$
4 · solve for the unknown	$x = \underline{\hspace{2cm}}$

Complete the table.

$$2(+1) + x + 4(-2) = 0 \rightarrow x = +6$$

$$\text{H: } +1, +1 \cdot \text{S: } +6 \cdot \text{O: } -2 \times 4 \cdot \text{check: } +2 + 6 - 8 = 0 \checkmark$$



Practice 1



calcium sulfite · a neutral compound

Power plants trap SO_2 from smoke with lime, which turns it into calcium sulfite. What is the oxidation number of S in CaSO_3 ?

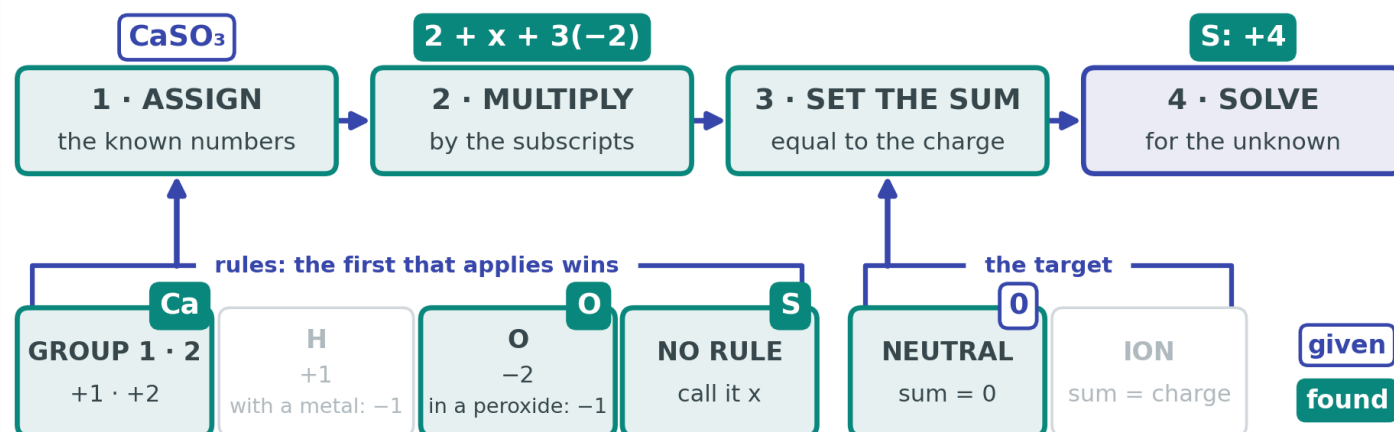
- A. +6
- B. -2
- C. +5
- D. +4
- E. 0

✓ Practice 1: answer D

$$(+2) + x + 3(-2) = 0 \rightarrow x = 0 - 2 + 6 = +4 \text{ (answer D)}$$

Ca: +2 (Group 2) · S: +4 · O: -2×3 · check: $+2 + 4 - 6 = 0$ ✓

🧠 **Sense check:** S is +4 in both SO_2 and CaSO_3 . Trapping the gas changed its partners, not its number. ✓



⚠ Where this goes wrong

Counting the -2 once for the whole formula. In CO_3^{2-} , $x + (-2) = -2$ gives $x = 0$. The -2 belongs to each O atom: three of them contribute $3(-2) = -6$, and $x + (-6) = -2$ gives $x = +4$.

Putting the unknown on the negative side. In ClO_3^- , writing Cl as -5 sums to $-5 - 6 = -11$, not -1 . Oxygen already holds the negative numbers; Cl balances them from the positive side: $-1 + 6 = +5$.

Giving H $+1$ next to a metal. In NaH, that reads $+1 + 1 = +2$, not 0. Sodium's rule sits higher on the list, so Na is $+1$ and H takes -1 : $+1 - 1 = 0$ ✓.

Giving O -2 in a peroxide. In H_2O_2 , that reads $2(+1) + 2(-2) = -2$, not 0. Peroxide oxygen is -1 : $2(+1) + 2(-1) = 0$ ✓.



Worked example 3: SO_4^{2-}

Step 1 · Assign the known numbers



O: -2 each (rule) · S: x · wanted: the oxidation number of S

The sulfate ion appears in fertilizer and hard water. A common first attempt sets the sum to zero. Test it.

✓ Worked example 3: testing the first attempt



O: -2 each (rule) · S: x · wanted: the oxidation number of S

A common first attempt

$$x + 4(-2) = 0 \rightarrow x = +8$$

sum: $+8 - 8 = 0$ ✗, but SO_4^{2-} is not neutral

A sum of zero describes a neutral species. This ion carries $2-$.

 **Sense check:** The check must land on the charge written on the species: -2 , not 0 .

✓ Worked example 3: solution



O: -2 each (rule) · S: x · wanted: the oxidation number of S

Step 2 · Multiply by the subscripts

Step 3 · Set the sum equal to the charge

$$x + 4(-2) = -2$$

S: x · O: -2 each, 4 atoms $\rightarrow -8$ · ion $\rightarrow$ sum -2 , not 0

Step 4 · Solve for the unknown

$$x = -2 + 8 = +6$$

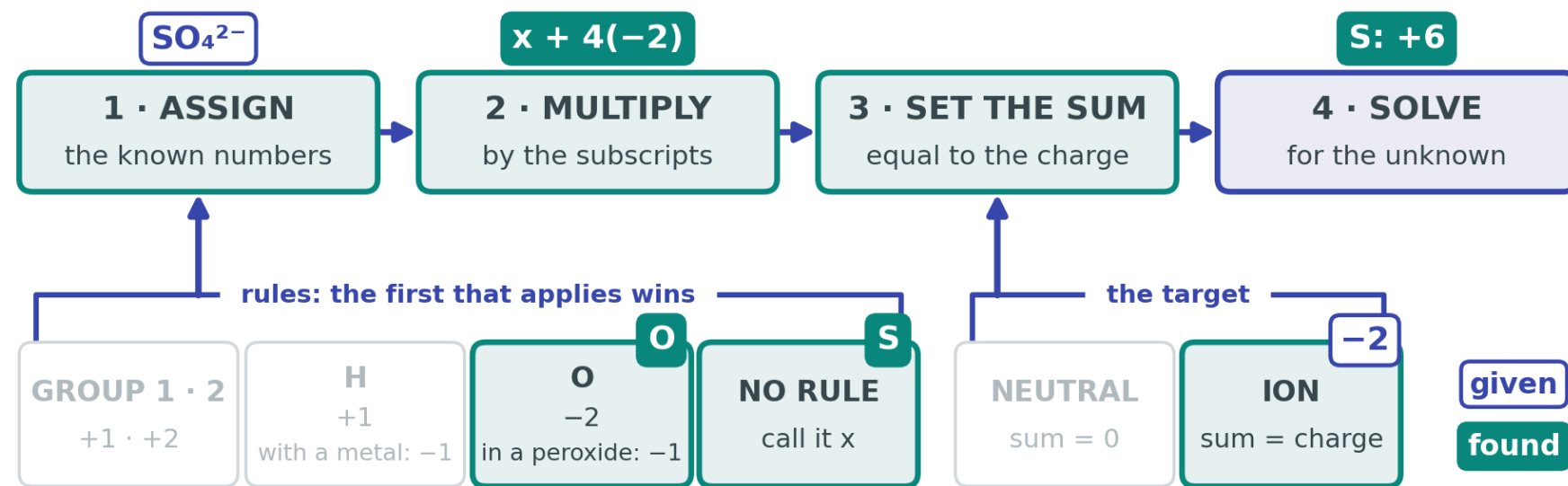
S: $+6$ · O: -2×4 · check: $+6 - 8 = -2$ ✓, the ion's charge

 **Sense check:** $+6 - 8 = -2$ ✓. The first attempt's $+8$ fails the same check: $+8 - 8 = 0 \neq -2$.

✓ Worked example 3: the route on the map



$$\text{found: } x + 4(-2) = -2 \cdot \text{ion} \rightarrow \text{sum } -2$$



Sense check: Same O card as SO₂. The ion card moved the target from 0 to -2. ✓



Practice 2



the oxalate ion · a polyatomic ion

Spinach and rhubarb leaves hold the oxalate ion. What is the oxidation number of each C atom in $\text{C}_2\text{O}_4^{2-}$?

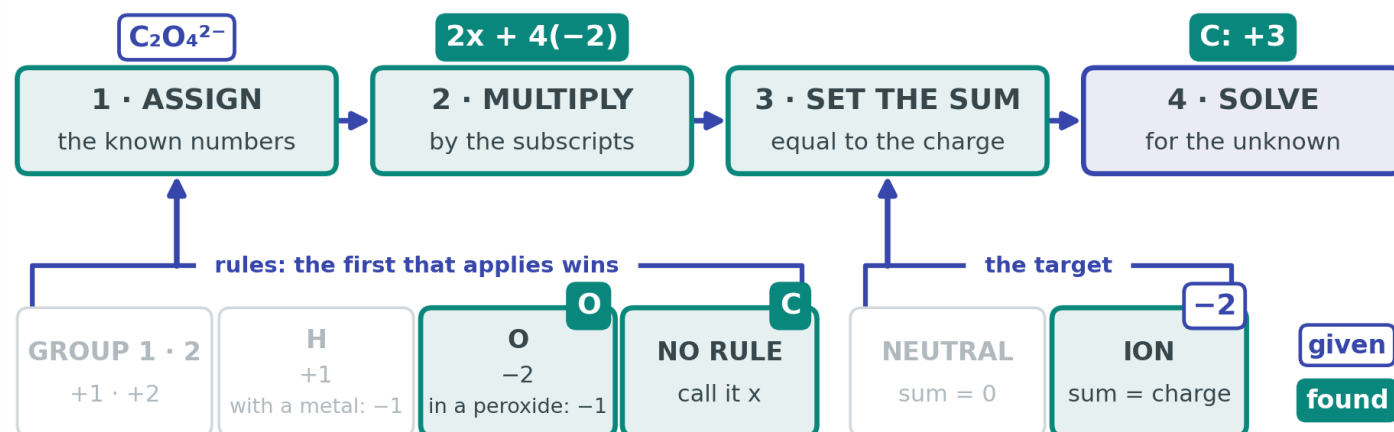
- A. +3
- B. +4
- C. +5
- D. +6

✓ Practice 2: answer A

$$2x + 4(-2) = -2 \rightarrow 2x = -2 + 8 = 6 \rightarrow x = +3 \text{ (answer A)}$$

C: +3 each, 2 atoms $\rightarrow +6$ · O: $-2 \times 4 \rightarrow -8$ · check: $+6 - 8 = -2$ ✓, the ion's charge

🧠 **Sense check:** Two alike C atoms share +6 equally: +3 each. The sum lands on the ion's charge, -2. ✓





Practice 3



ammonia, a neutral molecule · ammonium, a polyatomic ion

Fertilizer plants turn ammonia into ammonium salts. Compare the oxidation number of N in the two species. Which statement is correct?

- A. Different: -3 in NH_3 but -4 in NH_4^+ .
- B. The same in both: -3 . The sums differ, 0 and $+1$, yet both solve to -3 .
- C. In NH_4^+ , N simply carries the ion's charge, $+1$; only in NH_3 is it -3 .
- D. The same in both: $+3$, since the central atom takes the positive side.

✓ Practice 3: answer B

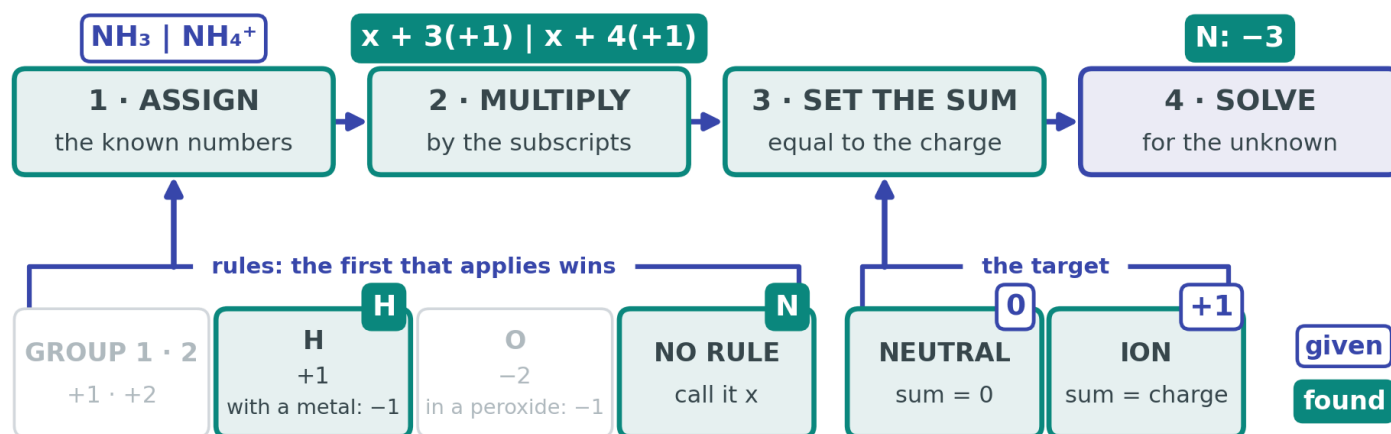
$$\text{NH}_3: x + 3(+1) = 0 \rightarrow x = 0 - 3 = -3$$

H: +1 (rule) · a neutral molecule: the sum's target is 0 · check: $-3 + 3 = 0$ ✓

$$\text{NH}_4^+: x + 4(+1) = +1 \rightarrow x = 1 - 4 = -3 \text{ (answer B)}$$

a polyatomic ion: the sum's target is its charge, +1 · check: $-3 + 4 = +1$ ✓

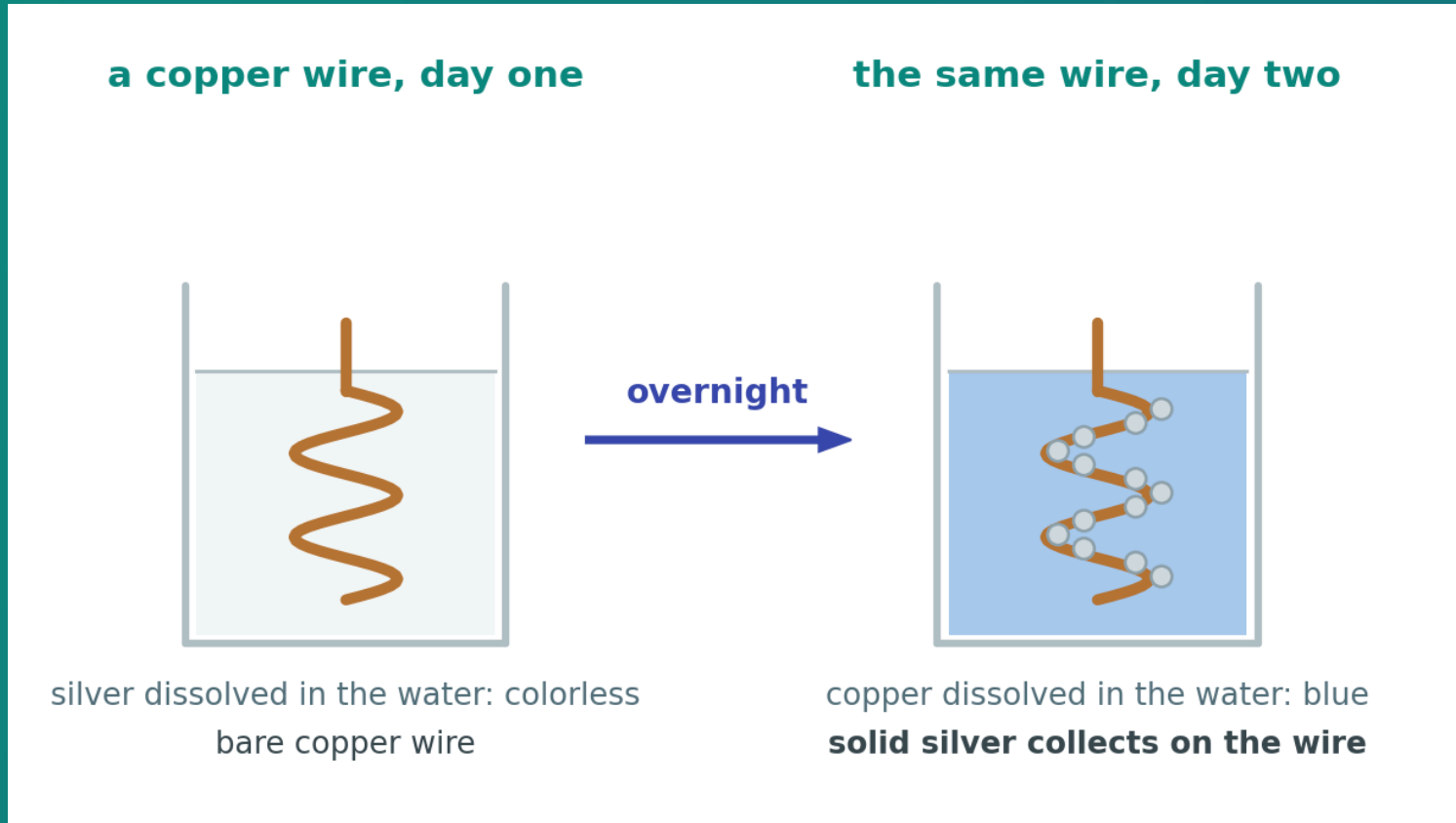
🧠 **Sense check:** Different targets, same rules, same answer. The extra H in NH_4^+ brought its own +1, and the ion's charge absorbs it.



8 · Oxidizing & Reducing Agents

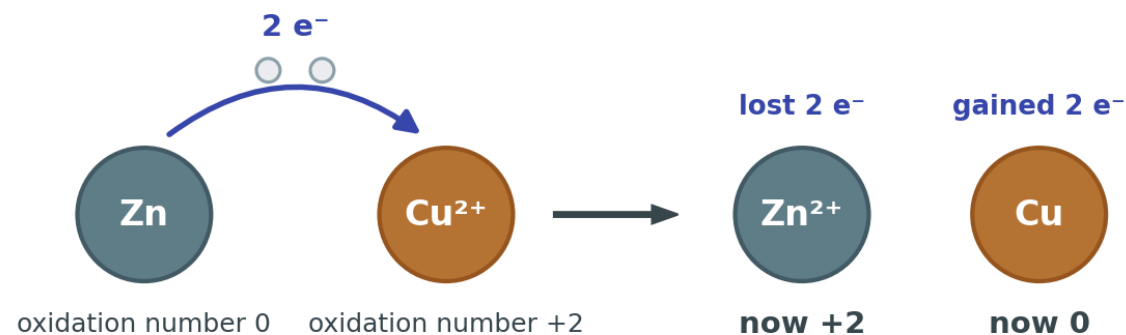
Decide from oxidation-number changes whether a reaction is redox, tell what is oxidized and what is reduced, and name the oxidizing and reducing agents.

Silver grows on copper



A copper wire stands overnight in dissolved silver. Solid silver collects on the wire; the liquid turns blue with dissolved copper. The two metals have traded places.

Electrons never vanish



electrons lost by Zn: 2 · electrons gained by Cu: 2

lost = gained: no electron vanishes ✓

Some reactions move electrons from one substance to another. Every electron one atom loses, another atom gains. Oxidation numbers make the transfer visible: they change only where electrons leave or arrive.

Ion charges are oxidation numbers



sodium metal + chlorine gas $\rightarrow$ table salt, made of Na^+ and Cl^- ions

Na: $0 \rightarrow +1$, gave 1 e^- · Cl: $0 \rightarrow -1$, took 1 e^-

2Na give 2 e^- · 2Cl take 2 e^- · NaCl : $(+1) + (-1) = 0$

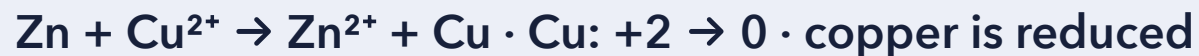
An element by itself counts 0. A simple ion's charge is its oxidation number. Forming an ion moves electrons: sodium gave one away, chlorine took it.

Oxidation and reduction

oxidation = the oxidation number increases: electrons lost



reduction = the oxidation number decreases: electrons gained



Assign oxidation numbers to both sides and compare. Electrons are negative, so losing them pushes the number up and gaining them pulls it down. OIL RIG: Oxidation Is Loss, Reduction Is Gain.

Naming the agents



Zn: $0 \rightarrow +2$ · Cu: $+2 \rightarrow 0$

what happens to it

Zn is oxidized

$0 \rightarrow +2$ · gives 2e^- away

what it is called

oxidizing agent

it takes electrons:
it oxidizes its partner

Cu^{2+} is reduced

$+2 \rightarrow 0$ · takes 2e^-

reducing agent

it gives electrons:
it reduces its partner

the agent name comes from what it does to its partner

Each agent is named for what it does to its partner, not for what happens to itself. Both agents are reactants: name the whole substance, not just the atom.

- 



Guided example: iron and sulfur

Step 1 • Assign oxidation numbers



Iron filings heated with sulfur glow red and leave a black solid, iron(II) sulfide. Fe and S alone count 0. In FeS, sulfur is the sulfide ion, -2 ; iron balances the formula: $x + (-2) = 0$ gives $+2$.

Which element is oxidized, which is reduced, and which reactant is each agent?

✓ Guided example: solution



Step 2 · Find the changes

Step 3 · Name the agents

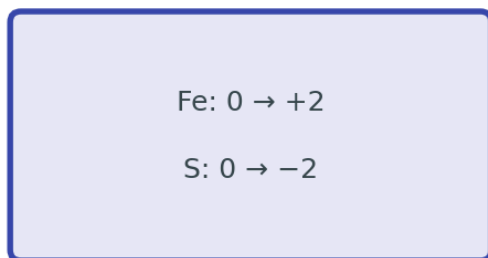
Fe gave its electrons to S: Fe is the **reducing agent**. S took them: S is the **oxidizing agent**.

🧠 **Sense check:** Electrons lost = gained: Fe lost 2, S gained 2. Each agent sits at the end of its own atom's lane. ✓

Step 1 · Assign oxidation numbers

Step 2 · Find the changes

Step 3 · Name the agents



number rises: **oxidized**
Fe: $0 \rightarrow +2$ · lost 2 e^-

number falls: **reduced**
S: $0 \rightarrow -2$ · gained 2 e^-

no number changes
no electrons moved

reducing agent: Fe

oxidizing agent: S

not redox
no agents

lost = gained: Fe lost 2 e^- · S gained 2 e^- ✓

Worked example 1: zinc in copper(II) sulfate

Step 1 • Assign oxidation numbers



A zinc strip in blue copper(II) sulfate solution darkens with copper, a single-displacement reaction. Elements by themselves count 0; O is -2 ; sulfate's sum, $x + 4(-2) = -2$, gives S $+6$; Cu balances its neutral formula at $+2$.

Which element is oxidized, and which is reduced?

✓ Worked example 1: solution



Step 2 · Find the changes

The unchanged atoms

 **Sense check:** Electrons lost = electrons gained: Zn lost 2, Cu gained 2. A reaction cannot lose electrons without something gaining them.

Step 1 · Assign oxidation numbers

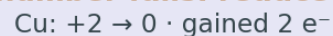


Step 2 · Find the changes

number rises: oxidized



number falls: reduced



no number changes

no electrons moved

Step 3 · Name the agents

reducing agent

its reactant gave the electrons

oxidizing agent

its reactant took the electrons

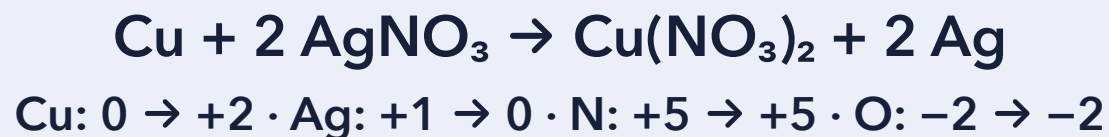
not redox

no agents

lost = gained: Zn lost 2 e⁻ · Cu gained 2 e⁻ ✓ · S and O unchanged

Worked example 2: copper in silver nitrate

Step 1 • Assign oxidation numbers



A copper wire in silver nitrate solution grows silver crystals while the liquid turns blue. Elements count 0; O is -2 ; nitrate's sum, $x + 3(-2) = -1$, gives N $+5$; Ag and Cu balance their formulas at $+1$ and $+2$.

Name both agents. A common first answer: copper is oxidized, so copper is the oxidizing agent. Test it against the definitions.

✓ Worked example 2: solution



Step 2 · Find the changes

A common first answer

Step 3 · Name the agents

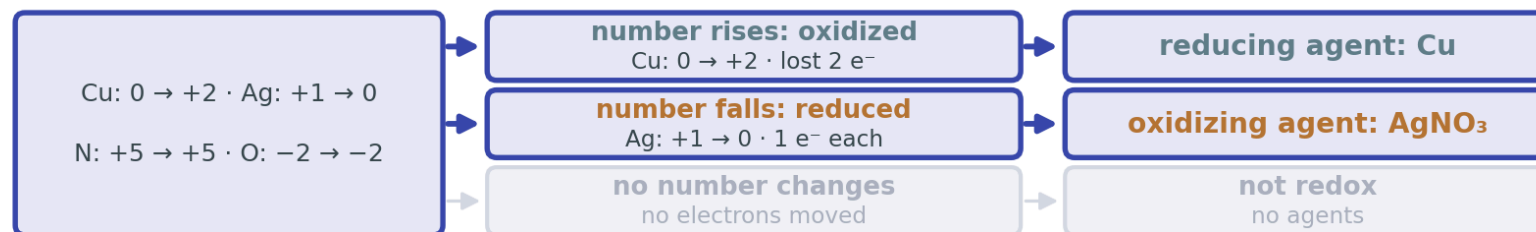
agent	substance	evidence
oxidizing agent	AgNO_3	contains Ag, reduced $+1 \rightarrow 0$
reducing agent	Cu	oxidized $0 \rightarrow +2$

 **Sense check:** The oxidized lane ends at the reducing agent: oxidized copper is the reducing agent. ✓

Step 1 · Assign oxidation numbers

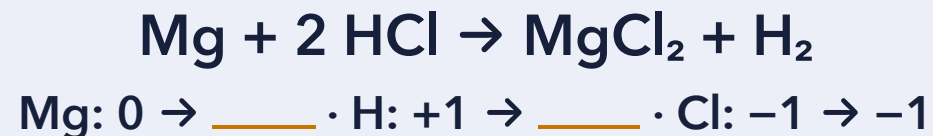
Step 2 · Find the changes

Step 3 · Name the agents



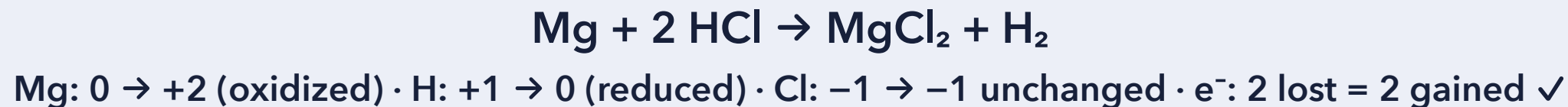
lost = gained: Cu lost 2e^- · $2 \text{Ag} \times 1 \text{e}^- = 2 \text{e}^-$ gained ✓

Your turn: magnesium in hydrochloric acid



question	answer
oxidized / reduced	_____
oxidizing agent	_____
reducing agent	_____

Complete the two changes, then name both agents.



HCl contains the reduced H: the **oxidizing agent**. Mg is oxidized: the **reducing agent**.

Where this goes wrong

Swapping the two agents. In $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$, calling Zn the oxidizing agent and CuSO_4 the reducing agent reverses both names. CuSO_4 takes zinc's electrons, and the reactant that takes electrons does the oxidizing. No metal is required for that job: in $\text{Cl}_2 + 2 \text{KBr} \rightarrow 2 \text{KCl} + \text{Br}_2$, the bare nonmetal Cl_2 takes bromide's electrons and is the oxidizing agent.

Calling the oxidized substance the oxidizing agent. Being oxidized means giving electrons to the partner, and giving electrons is what reduces the partner. The oxidized substance is always the **reducing** agent.

Picking a spectator as the agent. In $\text{Cu} + 2 \text{AgNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + 2 \text{Ag}$, nitrate moves from silver to copper, yet N stays +5 and O stays -2. Changing partners moves no electrons. Agents contain atoms whose numbers change.



Practice 1



A nickel bar sits in green copper(II) chloride solution and slowly plates with copper. Which species is the oxidizing agent?

- A. Ni, because it is the species that gets oxidized.
- B. CuCl_2 , because it contains Cu, which is reduced from +2 to 0.
- C. The chloride ion, Cl^- : it changes partners during the reaction, so it drives the electron transfer.
- D. Ni is the oxidizing agent, and CuCl_2 is the reducing agent.

✓ Practice 1: answer B



Ni: $0 \rightarrow +2$ (oxidized) · Cu: $+2 \rightarrow 0$ (reduced) · Cl: $-1 \rightarrow -1$ unchanged

Cu: $+2 \rightarrow 0$ · reduced · its reactant is CuCl_2 = **the oxidizing agent** (answer B)

🧠 **Sense check:** Electrons lost = gained: Ni lost 2, Cu gained 2. ✓

Step 1 · Assign oxidation numbers

Ni: $0 \rightarrow +2$ · Cu: $+2 \rightarrow 0$
Cl: $-1 \rightarrow -1$

Step 2 · Find the changes

number rises: oxidized
Ni: $0 \rightarrow +2$ · lost $2 e^-$

number falls: reduced
Cu: $+2 \rightarrow 0$ · gained $2 e^-$

no number changes
no electrons moved

Step 3 · Name the agents

reducing agent: Ni

oxidizing agent: CuCl_2

not redox
no agents

lost = gained: Ni lost $2 e^-$ · Cu gained $2 e^-$ ✓ · Cl unchanged



Worked example 3: testing whether a reaction is redox



Two colorless solutions mix and a bright yellow solid, PbI_2 , appears. Every formula changes partners.

Assign oxidation numbers to both sides: is any element oxidized or reduced?

✓ Worked example 3: solution

Step 1 · Assign oxidation numbers



Pb: +2 → +2 · N: +5 → +5 · O: -2 → -2 · K: +1 → +1 · I: -1 → -1

Step 2 · Find the changes

 **Sense check:** Electrons lost = gained: 0 = 0. A yellow precipitate is a real reaction, but ions that keep their numbers have only changed partners.

Step 1 · Assign oxidation numbers

Pb: +2 → +2 · K: +1 → +1
N: +5 → +5 · O: -2 → -2
I: -1 → -1

Step 2 · Find the changes

number rises: **oxidized**
electrons lost

number falls: **reduced**
electrons gained

no number changes
every number the same, both sides

Step 3 · Name the agents

reducing agent
its reactant gave the electrons

oxidizing agent
its reactant took the electrons

not redox: no agents

lost = gained: 0 = 0 · the ions only changed partners

Half-reactions



the electron is a product: X lost it · LEO, Lose Electrons: Oxidation



the electron is a reactant: Y^+ gained it · GER, Gain Electrons: Reduction

Any redox reaction splits into two half-reactions, the loss and the gain written separately. Each half must balance atoms and total charge. LEO GER and OIL RIG state the same rule.

Worked example 4: splitting a reaction into half-reactions



Zn: $0 \rightarrow +2$ · Cu: $+2 \rightarrow 0$ · wanted: both half-reactions, classified

A zinc strip in copper(II) solution, written as a net ionic equation. Split the reaction into its two half-reactions, classify each, and check that the electron counts match.

✓ Worked example 4: solution



The oxidation half-reaction



The reduction half-reaction



 **Sense check:** 2e^{-} leave the oxidation half and 2e^{-} enter the reduction half: lost = gained. Adding the halves cancels the electrons and rebuilds the equation.



Practice 2



Chlorine bubbled through pale green iron(II) chloride solution turns it yellow-brown. Which reactant is the reducing agent, and how many electrons change hands in the equation as written?

- A. Cl_2 is the reducing agent; 2 electrons change hands.
- B. FeCl_2 is the reducing agent; 4 electrons change hands.
- C. FeCl_2 is the reducing agent; 1 electron changes hands.
- D. FeCl_2 is the reducing agent; 2 electrons change hands.

✓ Practice 2: answer D



Fe: $+2 \rightarrow +3$ (oxidized) · Cl from Cl_2 : $0 \rightarrow -1$ (reduced) · Cl already in FeCl_2 : $-1 \rightarrow -1$

$2 \text{Fe} \times 1 \text{e}^- \text{ lost} = 2 \text{e}^- = 2 \text{Cl} \times 1 \text{e}^- \text{ gained}$

· FeCl_2 gives them: the reducing agent (answer D)

🧠 **Sense check:** Chlorine stands on both sides, yet only the two Cl atoms from Cl_2 change. Lost = gained: $2 = 2$ ✓.

Step 1 · Assign oxidation numbers

Fe: $+2 \rightarrow +3$
Cl from Cl_2 : $0 \rightarrow -1$
Cl in FeCl_2 : $-1 \rightarrow -1$

Step 2 · Find the changes

number rises: **oxidized**

Fe: $+2 \rightarrow +3$ · $2 \text{Fe} \times 1 \text{e}^-$

number falls: **reduced**

Cl: $0 \rightarrow -1$ · $2 \text{Cl} \times 1 \text{e}^-$

no number changes

no electrons moved

Step 3 · Name the agents

reducing agent: FeCl_2

oxidizing agent: Cl_2

not redox

no agents

lost = gained: $2 \text{e}^- = 2 \text{e}^-$ ✓ · 2 electrons change hands



Practice 3



Aluminum foil in silver nitrate solution slowly coats with silver. Which half-reaction belongs to the reducing agent?

- A. $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$
- B. $\text{Al} + 3 \text{e}^- \rightarrow \text{Al}^{3+}$
- C. $\text{Al} \rightarrow \text{Al}^{3+} + \text{e}^-$
- D. $\text{Al}^{3+} + 3 \text{e}^- \rightarrow \text{Al}$
- E. $\text{Al} \rightarrow \text{Al}^{3+} + 3 \text{e}^-$

✓ Practice 3: answer E



Al: $0 \rightarrow +3$ (oxidized) · Ag: $+1 \rightarrow 0$ (reduced) · N: $+5 \rightarrow +5$ · O: $-2 \rightarrow -2$

Al: $0 \rightarrow +3$ · oxidized: the reducing agent · 3e^- lost, a product

= $\text{Al} \rightarrow \text{Al}^{3+} + 3 \text{e}^-$ (answer E)

🧠 **Sense check:** Electrons lost = gained: $1 \text{Al} \times 3 \text{e}^- = 3 \text{Ag} \times 1 \text{e}^- = 3$. Charge: left 0, right $(3+) + 3(1-) = 0$ ✓

Step 1 · Assign oxidation numbers

Al: $0 \rightarrow +3$ · Ag: $+1 \rightarrow 0$
N: $+5 \rightarrow +5$ · O: $-2 \rightarrow -2$

Step 2 · Find the changes

number rises: oxidized
 $\text{Al} \rightarrow \text{Al}^{3+} + 3 \text{e}^-$ · e^- a product

number falls: reduced
 $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$ · $3 \text{Ag} \times 1 \text{e}^-$

no number changes
no electrons moved

Step 3 · Name the agents

reducing agent: Al

oxidizing agent: AgNO_3

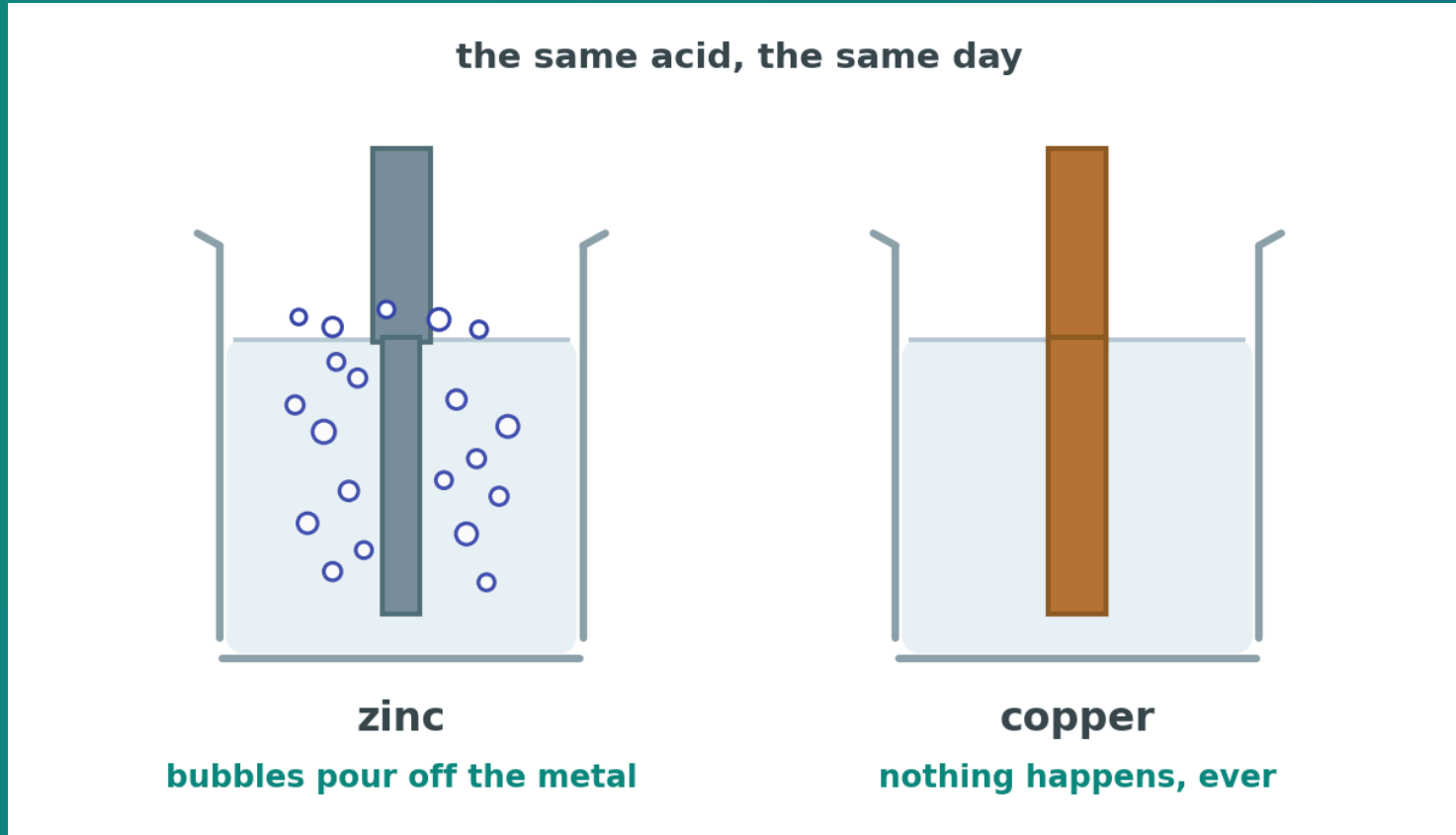
not redox
no agents

lost = gained: $1 \text{Al} \times 3 \text{e}^- = 3 \text{Ag} \times 1 \text{e}^- = 3$ ✓

9 • The Activity Series

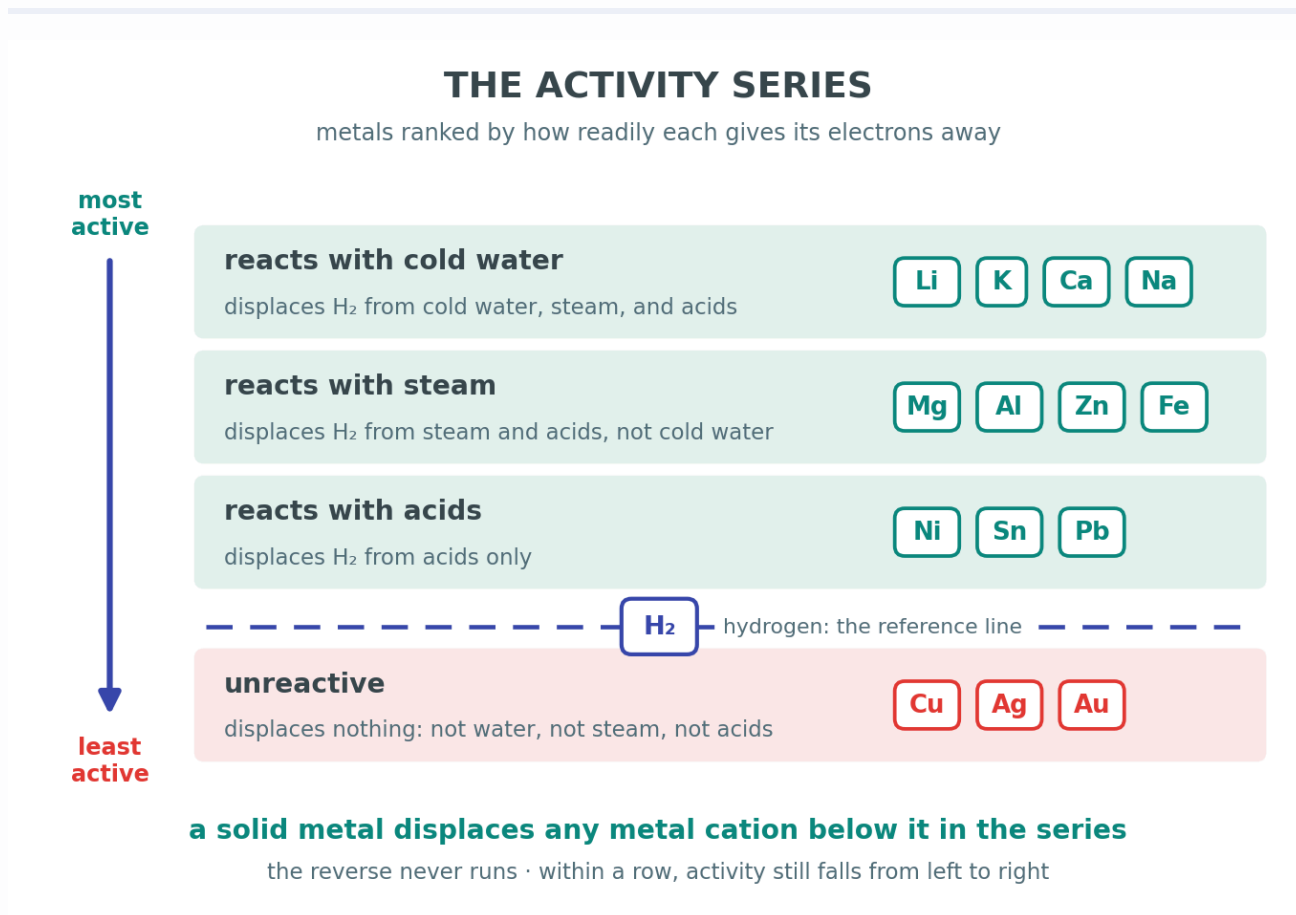
Use the activity series to predict whether a single-displacement reaction runs, write the products when it does, and write no reaction when it does not.

Two metals, one acid



Zinc fizzes furiously in hydrochloric acid; copper sits in the same acid untouched, forever. Something ranks these metals.

The activity series



Four tiers, ranked by how readily each metal gives electrons away: displace any cation below you, never one above.

Two guaranteed product patterns



any metal above hydrogen pushes its electrons onto H^+



the cold-water tier takes its hydrogen from water itself

The series is a measured ranking of how readily each metal gives its electrons away. A metal above hydrogen runs its oxidation half-reaction against H^+ , and the products follow the pattern.



The method

1. **Find both metals in the series.** With acid or water, hydrogen is the second metal.
2. **Compare positions.** The solid metal must sit above its partner; otherwise write no reaction.
3. **Write the products from the pattern.**
4. **Balance the equation.**



Worked example 1: magnesium in hydrochloric acid

Step 1 · Find both metals in the series



Mg: the steam tier · H: supplied by the acid · wanted: products, balanced

A magnesium strip dropped into hydrochloric acid dissolves in a rush of bubbles. Predict the products and balance the equation.

✓ Worked example 1: solution



Mg: the steam tier · H: supplied by the acid

Step 2 · Compare positions

Step 3 · Write the products from the pattern



charge in the salt: $(2+) + 2(1-) = 0$ ✓

Step 4 · Balance the equation



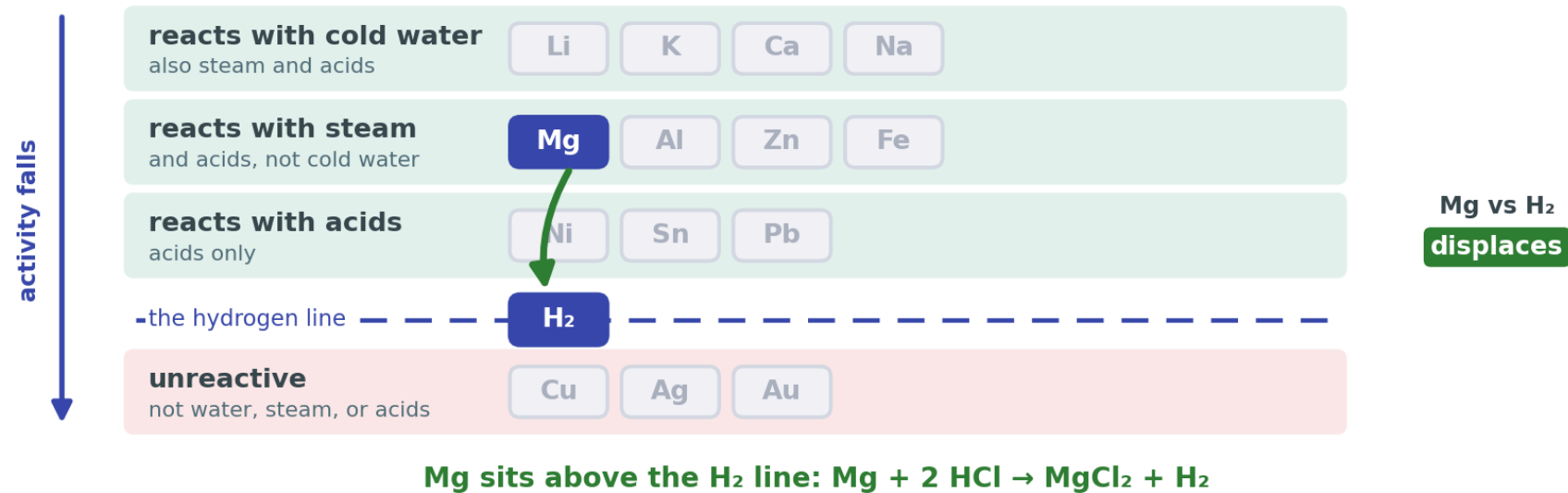
Mg: $1 = 1$ ✓ · H: $2 = 2$ ✓ · Cl: $2 = 2$ ✓

 **Sense check:** The bubbles are the H_2 the pattern promised. Mg loses 2 e^- ; two H^+ gain 1 each: lost = gained, $2 = 2$ ✓.

✓ Worked example 1: the route on the series



solid metal: Mg · partner: hydrogen, from the acid · found: the reaction runs



🧠 **Sense check:** With an acid, the partner is hydrogen. Mg sits above the H₂ line, so the reaction runs. ✓



Worked example 2: aluminum in nickel(II) nitrate

Step 1 • Find both metals in the series



Al: the steam tier • Ni: the acid tier, lower • wanted: products, balanced

Aluminum foil sits in green nickel(II) nitrate solution and slowly darkens with nickel. Predict the products and balance the equation.

✓ Worked example 2: solution



Al: the steam tier · Ni: the acid tier, lower

Step 2 · Compare positions

Step 3 · Write the products from the pattern



charge in the salt: $(3+) + 3(1-) = 0 \checkmark$

Step 4 · Balance the equation



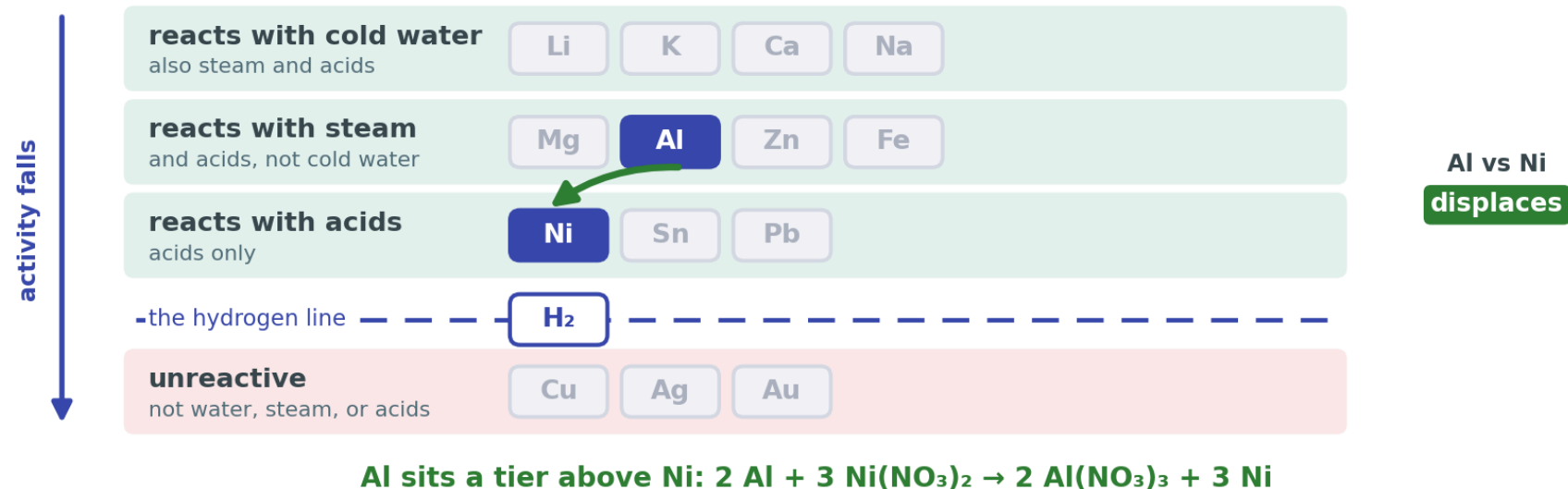
Al: $2 = 2 \checkmark$ · Ni: $3 = 3 \checkmark$ · N: $3(2) = 6 = 2(3) \checkmark$ · O: $3(6) = 18 = 2(9) \checkmark$

 **Sense check:** Electrons match too: 2 Al lose $3 e^-$ each, 6 total; 3 Ni^{2+} gain 2 each, 6 total. The nitrate ions only change partners.

✓ Worked example 2: the route on the series



solid metal: Al · partner: Ni, in the compound · found: the reaction runs



🧠 **Sense check:** Al sits in the steam tier, a full tier above Ni. The higher metal displaces the lower cation. ✓

Your turn: calcium in cold water



the pattern: active metal + water $\rightarrow$ metal hydroxide + H_2

step	work
1 · find both metals	Ca: the cold-water tier · H: from the water itself
2 · compare positions	Ca sits _____ hydrogen: the reaction runs
3 · products from the pattern	Ca(OH)_2 + _____
4 · balance	$\text{Ca} + \text{_____ H}_2\text{O} \rightarrow \text{Ca(OH)}_2 + \text{H}_2$

Complete the table.



Ca: 1 = 1 ✓ · O: 2 = 2 ✓ · H: 2(2) = 4 = 2 + 2 ✓

⚠ Where this goes wrong

Running the series upward. Zinc displaces Cu^{2+} , so it is tempting to write the reverse: $\text{Cu} + \text{ZnSO}_4 \rightarrow \text{CuSO}_4 + \text{Zn}$. Copper sits below zinc and cannot hand its electrons up the list. The reverse of a working displacement never runs.

Never writing "no reaction." $\text{Ag} + \text{ZnCl}_2$ swaps neatly on paper into $\text{AgCl} + \text{Zn}$, and it never happens: silver sits in the unreactive tier, below zinc. When the solid metal lies below the dissolved cation, no reaction is the complete answer.



Practice 1



Zn: the steam tier · Pb: the acid tier, lower

A zinc strip hangs in lead(II) nitrate solution, and gray crystals grow on it. Which equation describes the change?

- A. $\text{Zn} + \text{Pb(NO}_3)_2 \rightarrow \text{Zn(NO}_3)_2 + \text{Pb}$: zinc, above lead, hands its electrons to Pb^{2+}
- B. No reaction: lead sits above zinc, so zinc cannot displace it
- C. $\text{Zn} + \text{Pb(NO}_3)_2 \rightarrow \text{ZnNO}_3 + \text{Pb}$: the displacement runs, with zinc as a 1+ ion
- D. Both directions run: $\text{Pb} + \text{Zn(NO}_3)_2 \rightarrow \text{Pb(NO}_3)_2 + \text{Zn}$ works just as well

✓ Practice 1: answer A



Zn: 1 = 1 ✓ · Pb: 1 = 1 ✓ · N: 2 = 2 ✓ · O: 6 = 6 ✓ · Zn above Pb: the reaction runs

B reverses the ranking: zinc stands in the steam tier and lead in the acid tier below it. C has the displacement right and the salt wrong: zinc forms a 2+ ion and nitrate is 1−, so the salt is $\text{Zn}(\text{NO}_3)_2$, $(2+) + 2(1-) = 0$. D runs the series upward: lead, the lower metal, never displaces Zn^{2+} .

 **Sense check:** Zn loses 2 e[−] and Pb²⁺ gains 2: lost = gained ✓. The gray crystals are lead metal leaving the solution.



Worked example 3: copper in lithium nitrate

Step 1 · Find both metals in the series



Cu: the unreactive tier · Li: the cold-water tier, at the very top

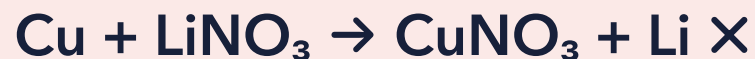
A copper wire stands in lithium nitrate solution. A common first attempt swaps the partners into $\text{CuNO}_3 + \text{Li}$. Test it against the series.

✓ Worked example 3: solution



Cu: the unreactive tier · Li: the cold-water tier, at the very top

A common first attempt



this asks copper to push its electrons onto Li^+ , three tiers up

Step 2 · Compare positions



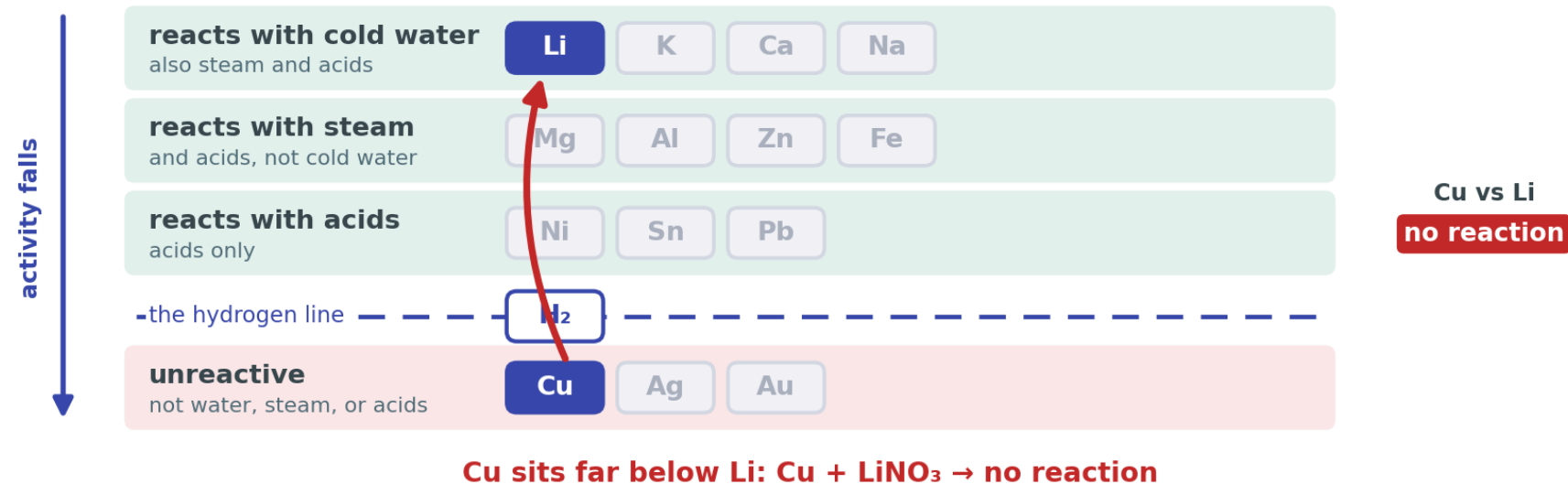
solid metal below the dissolved cation: no electron hand-off

 **Sense check:** The wire can stand there for years. A swap that looks fine on paper still needs the electron transfer to run downhill in the series.

✓ Worked example 3: the route on the series



solid metal: Cu · partner: Li, in the compound · found: no reaction



🧠 **Sense check:** The arrow points up the series. Copper sits far below lithium, so nothing happens. ✓



Practice 2



observed: an iron nail in nickel(II) sulfate solution darkens with nickel

A classmate reasons that swapping the roles must work as well, and writes $\text{Ni} + \text{FeSO}_4 \rightarrow \text{NiSO}_4 + \text{Fe}$. Which verdict on that claim is correct?

- A. It stands: every atom count matches, so the reverse runs.
- B. It fails: Ni sits below Fe, so it cannot displace Fe^{2+} . $\text{Ni} + \text{FeSO}_4$ gives no reaction.
- C. It stands: Ni and Fe both sit above hydrogen, so each can displace the other.
- D. It stands, slowly: the reverse runs, just more slowly, since Ni is the less active metal.

✓ Practice 2: answer B



Fe: the steam tier · Ni: the acid tier, lower · the solid metal sits below the dissolved cation

A treats balance as permission: Ni 1 = 1, Fe 1 = 1, S 1 = 1, O 4 = 4, yet a balanced equation can describe a reaction that never happens. C compared each metal with hydrogen; the partner here is Fe^{2+} , which sits above Ni. D: speed never moves a metal up the series; the ranking decides whether electrons flow at all.

 **Sense check:** The forward reaction is the evidence: Fe gave its electrons to Ni^{2+} . The same ranking forbids Ni from handing them back.

Summary of reaction types

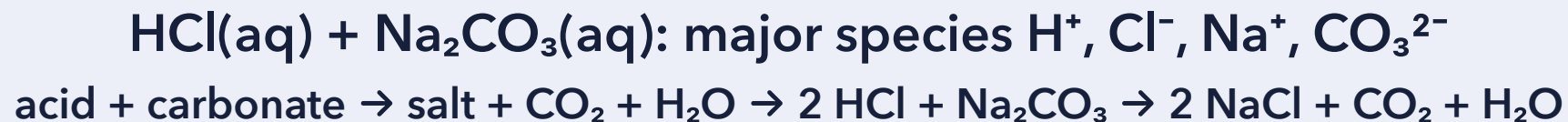
SUMMARY OF REACTION TYPES

you mix	decision tool	products
two dissolved salts	solubility rules	insoluble salt (s) + soluble salt
acid + hydroxide	runs as written	salt + water
acid + carbonate or bicarbonate	runs as written	salt + CO ₂ + H ₂ O
metal + acid or water	activity series	salt or hydroxide + H ₂
metal + metal cation (aq)	activity series	new salt + displaced metal
hydrocarbon + O ₂	runs as written	CO ₂ + H ₂ O

to predict: list the major species present, then find the matching row

Six reactant patterns cover nearly every product prediction. The solubility rules pick the precipitate; the activity series decides whether displacement runs.

Predicting products in aqueous solutions

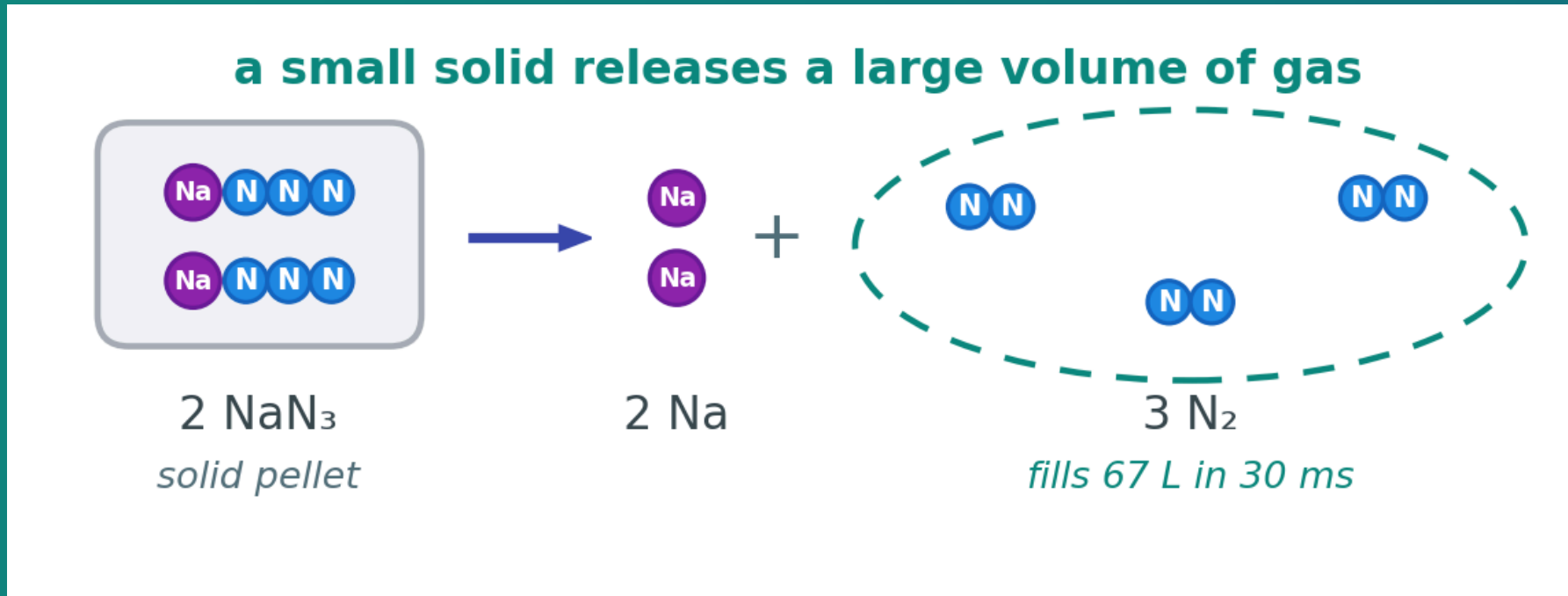


List the major species actually in the beaker. Match them to a reaction-type pattern; the first match names the products. No match on any pattern means no reaction.

10 • Mole Ratios

Use the coefficients of a balanced equation to convert moles of one substance into moles of any other.

The airbag problem



In a crash, an airbag pellet of sodium azide decomposes into nitrogen gas: **67 liters in 30 milliseconds**. Mole ratios determine how much solid to pack.

Why coefficients mean moles



2 : 5 : 4 : 2 molecules

2 : 5 : 4 : 2 dozen

2 : 5 : 4 : 2 mol, the same ratio at every scale

A balanced equation counts **molecules**. Scaling every amount by the same number keeps the ratio. Avogadro's number is one such multiplier: **coefficients count moles too**.

Two numbers appear in **2 H₂O**. Name each one's job.



which number may balancing change · which is part of the formula itself

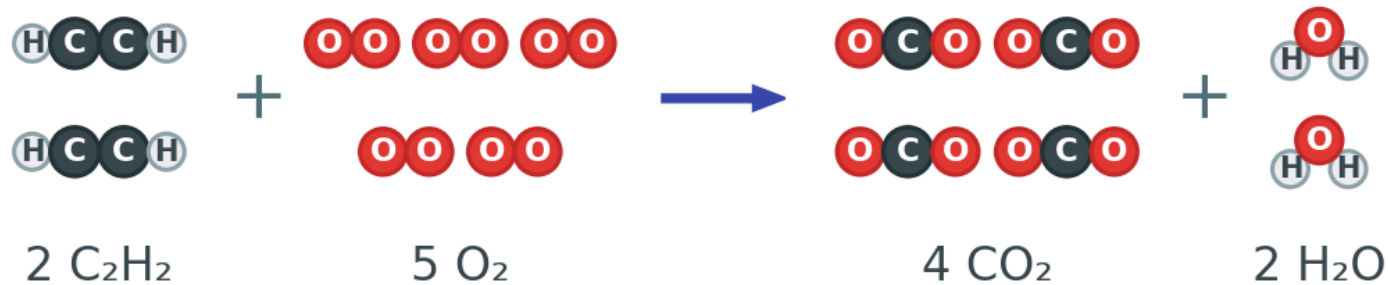


the coefficient 2 may change during balancing ✓ · changing the subscript ₂ would name a different substance ✗

Reading the coefficients



the coefficients count the molecules: $2 \cdot 5 \rightarrow 4 \cdot 2$



The coefficients count the molecules in the figure. The same numbers apply in **moles**. Like a recipe, the amounts scale together.

Coefficients become conversion factors



Any two coefficients form a **mole ratio**: a fraction that converts moles of one substance into moles of another.

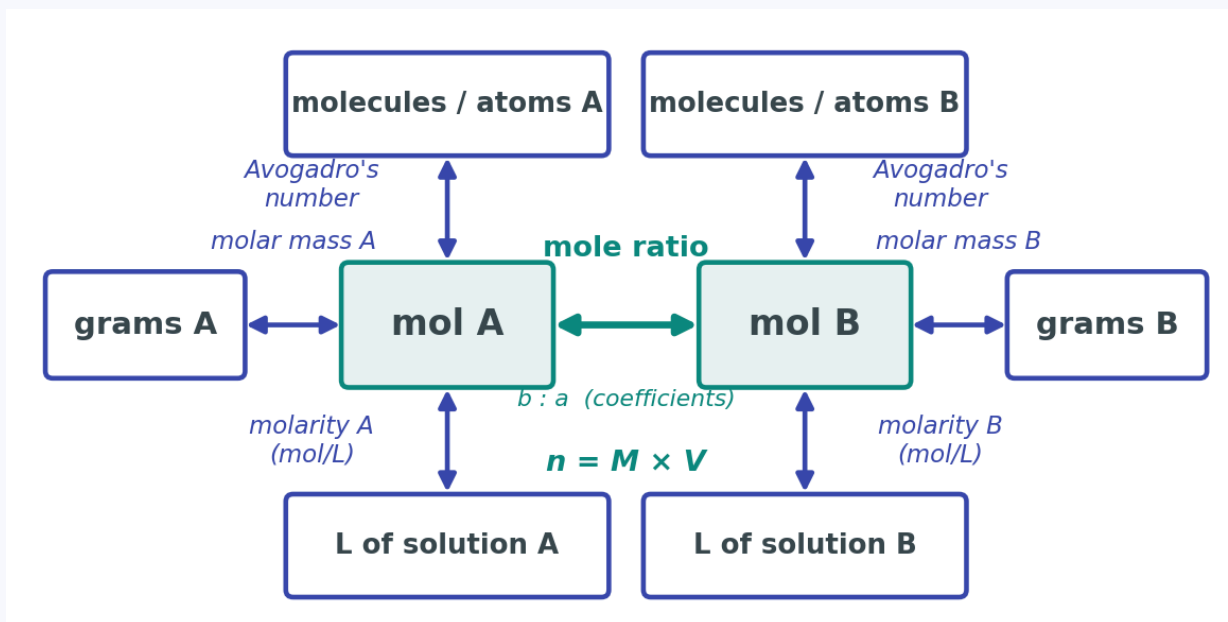
$$\frac{4 \text{ mol CO}_2}{5 \text{ mol O}_2} \text{ or } \frac{5 \text{ mol O}_2}{2 \text{ mol C}_2\text{H}_2} \text{ or any pair}$$

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

The method

The heart of every stoichiometry problem is a mole to mole conversion.

1. Start from a **balanced equation**.
2. Build the ratio **wanted over given**, so the given unit cancels.
3. **Multiply and cancel**; check the result against the coefficients.





Worked example 1

Step 1 • Start from a balanced equation



given: 4.5 mol C_2H_2 • wanted: mol CO_2

A welding torch burns **4.5 mol of C_2H_2** . How many moles of CO_2 form?

Set it up: which ratio cancels mol C_2H_2 ?

✓ Worked example 1: solution



given: 4.5 mol C₂H₂

Step 2 • Build the ratio, wanted over given

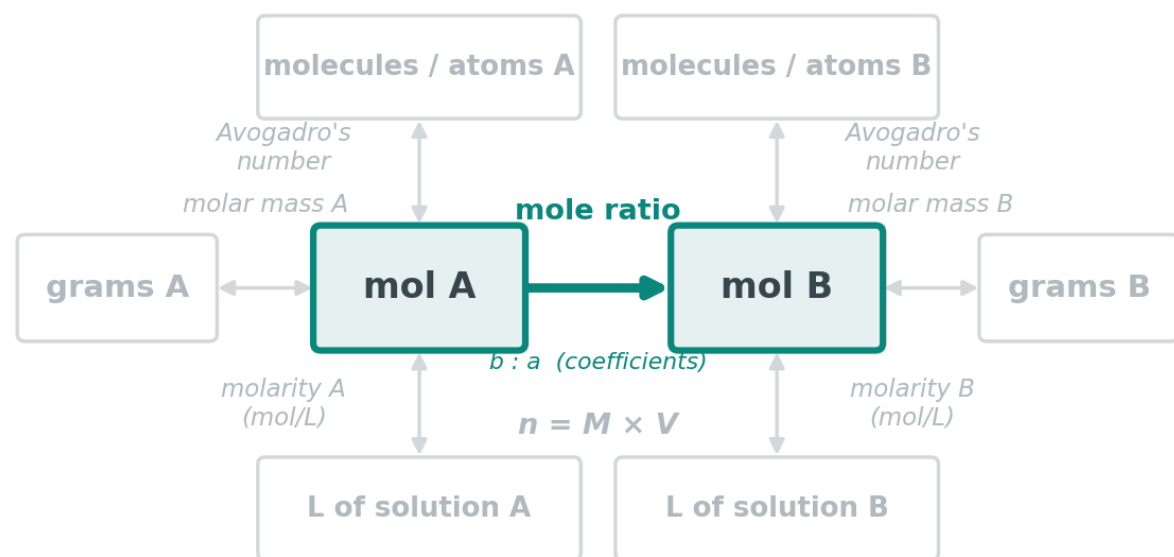
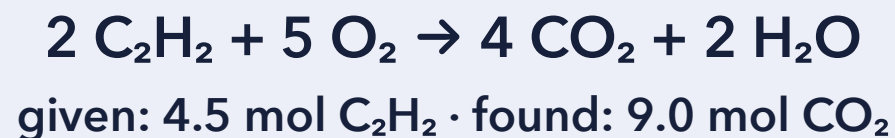
$$\frac{4 \text{ mol CO}_2}{2 \text{ mol C}_2\text{H}_2} \text{ cancels mol C}_2\text{H}_2 \checkmark \quad \frac{2 \text{ mol C}_2\text{H}_2}{4 \text{ mol CO}_2} \text{ cancels nothing } \times$$

Step 3 • Multiply and cancel

$$4.5 \text{ mol C}_2\text{H}_2 \times \frac{4 \text{ mol CO}_2}{2 \text{ mol C}_2\text{H}_2} = 9.0 \text{ mol CO}_2$$

🧠 **Sense check:** The coefficients make CO₂ double the C₂H₂ (4 vs 2), and 9.0 is double 4.5. ✓

✓ Worked example 1: the route on the map



🧠 **Sense check:** A is C₂H₂, B is CO₂. One arrow, one conversion factor: the mole ratio. ✓

Your turn: oxygen for the torch



given: 3.2 mol C_2H_2 · wanted: mol O_2

The torch consumes O_2 while it burns **3.2 mol of C_2H_2** :

$$3.2 \text{ mol } \cancel{\text{C}_2\text{H}_2} \times \frac{\text{ } \text{mol } \text{O}_2}{\text{ } \cancel{\text{mol } \text{C}_2\text{H}_2}} = \text{ } \text{mol } \text{O}_2$$

Complete the ratio so mol C_2H_2 cancels, then compute.

$$3.2 \text{ mol } \cancel{\text{C}_2\text{H}_2} \times \frac{5 \text{ mol } \text{O}_2}{2 \cancel{\text{mol } \text{C}_2\text{H}_2}} = \mathbf{8.0 \text{ mol } \text{O}_2}$$

 **Sense check:** O_2 's coefficient (5) is larger than C_2H_2 's (2), so the torch needs more O_2 than fuel: $8.0 > 3.2$. ✓

Where this goes wrong



Assuming 1:1. "6 mol $\text{H}_2 \rightarrow 6$ mol NH_3 ." The equation gives a 2 NH_3 : 3 H_2 ratio, so the answer is 4.0 mol.

Inverting the ratio. (3 mol H_2 / 2 mol NH_3) leaves units of mol H_2^2 /mol NH_3 . Nothing cancels. If the units do not cancel, the fraction is inverted.

Answering with the coefficient. The coefficient (2) is not the answer. It must be applied to the given amount.

 Practice 1

7.00 mol of O_2 react completely. How many moles of H_2O form?

- A. 7.00
- B. 14.0
- C. 3.50
- D. 2.00

✓ Practice 1: answer B



given: 7.00 mol O₂

$$7.00 \text{ mol } \cancel{\text{O}_2} \times \frac{2 \text{ mol H}_2\text{O}}{1 \cancel{\text{ mol O}_2}} = 14.0 \text{ mol H}_2\text{O (answer B)}$$

A assumed 1:1: $7.00 \times 1 = 7.00$. C inverted the ratio: $7.00 \times (1/2) = 3.50$. D answered with the coefficient, 2.

 **Sense check:** Water's coefficient is double O₂'s, so the answer is double the given. ✓

Worked example 2: moles of reactant needed

Step 1 • Start from a balanced equation

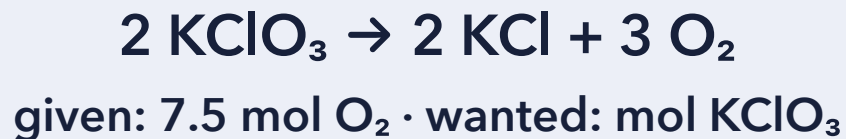


K: $2 = 2 \checkmark$ • Cl: $2 = 2 \checkmark$ • O: $6 = 6 \checkmark$ • given: 7.5 mol O_2 • wanted: mol KClO_3

Heating potassium chlorate releases oxygen gas, one design for emergency oxygen generators. A generator must deliver **7.5 mol of O_2** . How many moles of KClO_3 must it hold?

The given sits on the product side. The steps do not change.

✓ Worked example 2: solution



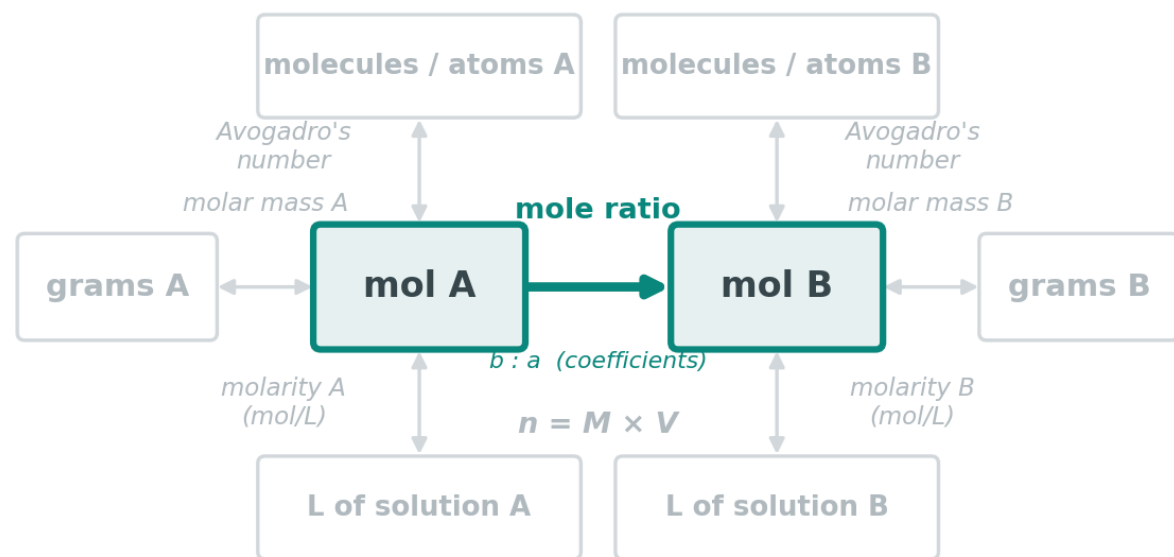
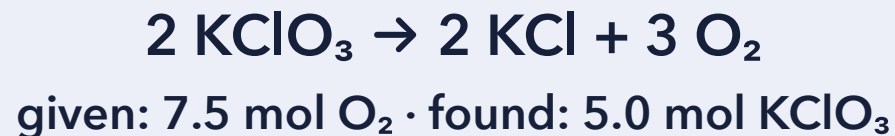
Step 2 · Build the ratio, wanted over given

Step 3 · Multiply and cancel

$$7.5 \text{ mol O}_2 \times \frac{2 \text{ mol KClO}_3}{3 \text{ mol O}_2} = 5.0 \text{ mol KClO}_3$$

 **Sense check:** Fewer moles of solid are packed than moles of gas delivered: 2 KClO₃ yield 3 O₂. The ratio converts in either direction across the equation. ✓

✓ Worked example 2: the route on the map



🧠 **Sense check:** A is the given, O₂, even though it is a product. B is KClO₃. The route is the same single arrow. ✓



Practice 3



Greenhouse growers burn butane to enrich the air with CO_2 , which speeds plant growth. How many moles of C_4H_{10} must burn to produce **26.0 mol of CO_2** ?

- A. 26.0
- B. 52.0
- C. 6.50
- D. 104

✓ Practice 3: answer C



given: 26.0 mol CO_2 · wanted: mol C_4H_{10}

$$26.0 \text{ mol } \cancel{\text{CO}_2} \times \frac{2 \text{ mol C}_4\text{H}_{10}}{8 \cancel{\text{ mol CO}_2}} = 6.50 \text{ mol C}_4\text{H}_{10} \text{ (answer C)}$$

A assumed a 1 : 1 ratio: $26.0 \times 1 = 26.0$. B multiplied by 2 without dividing by 8: $26.0 \times 2 = 52.0$. D inverted the ratio: $26.0 \times (8/2) = 104$.

 **Sense check:** The ratio 8 : 2 means four CO_2 per butane, so the fuel needed is $26.0 \div 4 = 6.50$ mol. ✓

 Practice 4

A refinery flare burns waste pentane. How many moles of gas in total, CO_2 plus H_2O vapor, leave the flare while it consumes **2.40 mol of O_2** ?

- A. 1.50
- B. 4.80
- C. 7.04
- D. 3.30

✓ Practice 4: answer D



given: 2.40 mol O₂ · wanted: total mol CO₂ + H₂O

$$2.40 \cancel{\text{mol O}_2} \times \frac{5 \text{ mol CO}_2}{8 \cancel{\text{mol O}_2}} = 1.50 \text{ mol CO}_2 \quad 2.40 \cancel{\text{mol O}_2} \times \frac{6 \text{ mol H}_2\text{O}}{8 \cancel{\text{mol O}_2}} = 1.80 \text{ mol H}_2\text{O}$$

$$1.50 \text{ mol CO}_2 + 1.80 \text{ mol H}_2\text{O} = \mathbf{3.30 \text{ mol gas (answer D)}}$$

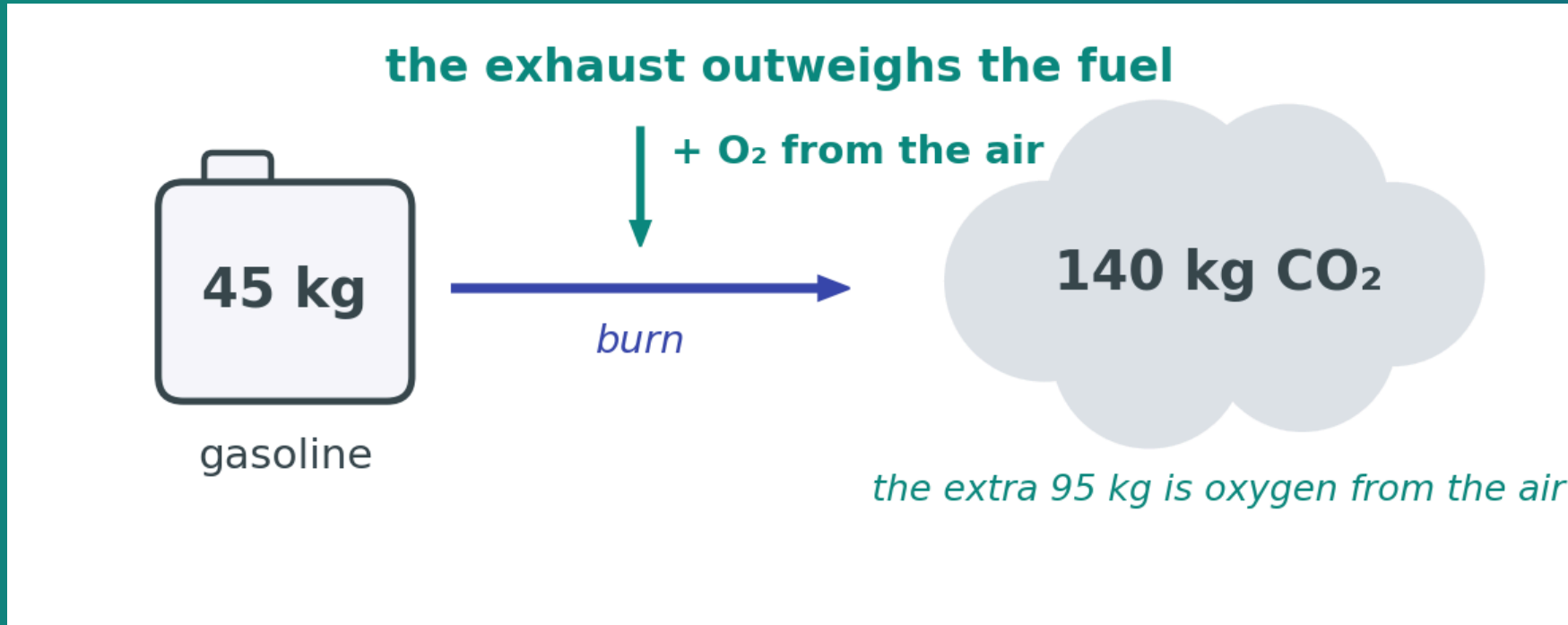
A stopped at the CO₂ alone: 1.50. B assumed 1:1 for each product: 2.40 + 2.40 = 4.80. C inverted both ratios: 2.40 × (8/5) + 2.40 × (8/6) = 3.84 + 3.20 = 7.04.

🧠 **Sense check:** 8 O₂ in, 5 + 6 = 11 gas molecules out: 3.30 > 2.40 ✓

11 · Mass-to-Mass Stoichiometry

Convert a given mass of one substance into the mass or molecule count of another by converting grams to moles, crossing substances with the mole ratio, and converting back.

Heavier than the fuel



One 45-kg tank of gasoline emits about **140 kg of CO₂**: triple the fuel's mass. The extra mass comes from oxygen in the air, and stoichiometry predicts it exactly.

Equations count particles, not grams

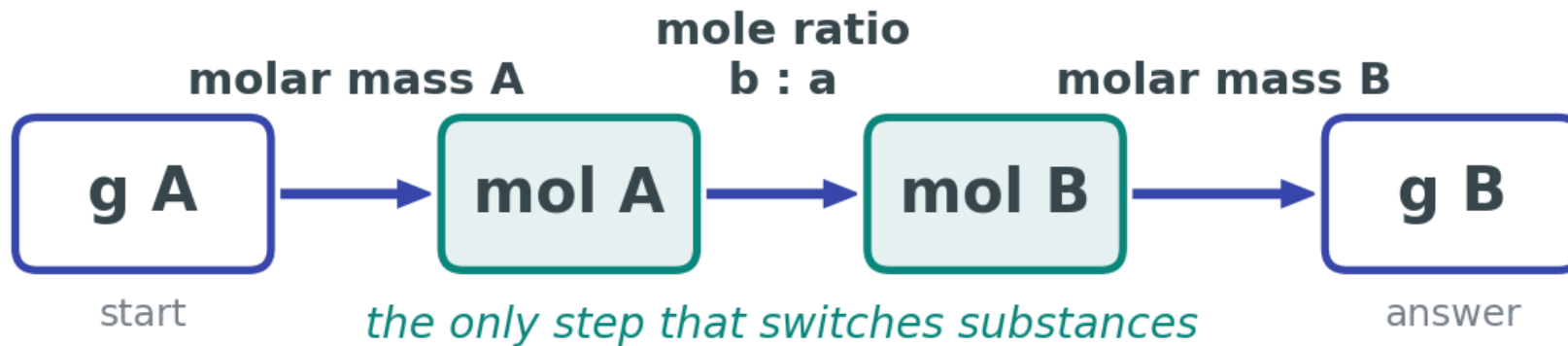


moles = 1 glucose : 6 CO₂ · grams = 180.16 : 264.06, not 1 : 6

A balanced equation relates **counts** (moles), never masses. A mole of glucose weighs four times a mole of CO₂, so a 1:6 mole ratio is **not** a 1:6 gram ratio.

Convert to moles, then convert back

Convert the given mass to **moles**, relate moles with the equation, and convert back to mass at the end.



Three conversion factors, one setup

Molar mass converts at each end; the mole ratio is the only factor that **switches substances**. Chain them so each unit cancels the one before:

$$\text{g A} \times \frac{1 \text{ mol A}}{(\text{molar mass A}) \text{ g A}} \times \frac{b \text{ mol B}}{a \text{ mol A}} \times \frac{(\text{molar mass B}) \text{ g B}}{1 \text{ mol B}} = \text{g B}$$



The method

1. **Grams → moles:** convert the given mass with its own molar mass.
2. **Moles → moles:** cross substances with the mole ratio. No other step can.
3. **Moles → grams or molecules:** convert out with the target's molar mass, or Avogadro's number.



Worked example 1: moles of product from a mass

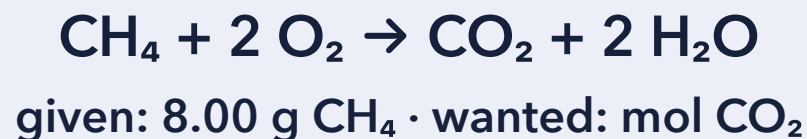


given: 8.00 g CH₄ · wanted: mol CO₂

A gas burner consumes **8.00 g of CH₄**. How many **moles** of CO₂ form? (CH₄ 16.04 g/mol, CO₂ 44.01 g/mol)

Write the route first: g CH₄ → mol CH₄ → mol CO₂. Stop at moles.

✓ Worked example 1: solution



Step 1 · Grams → moles

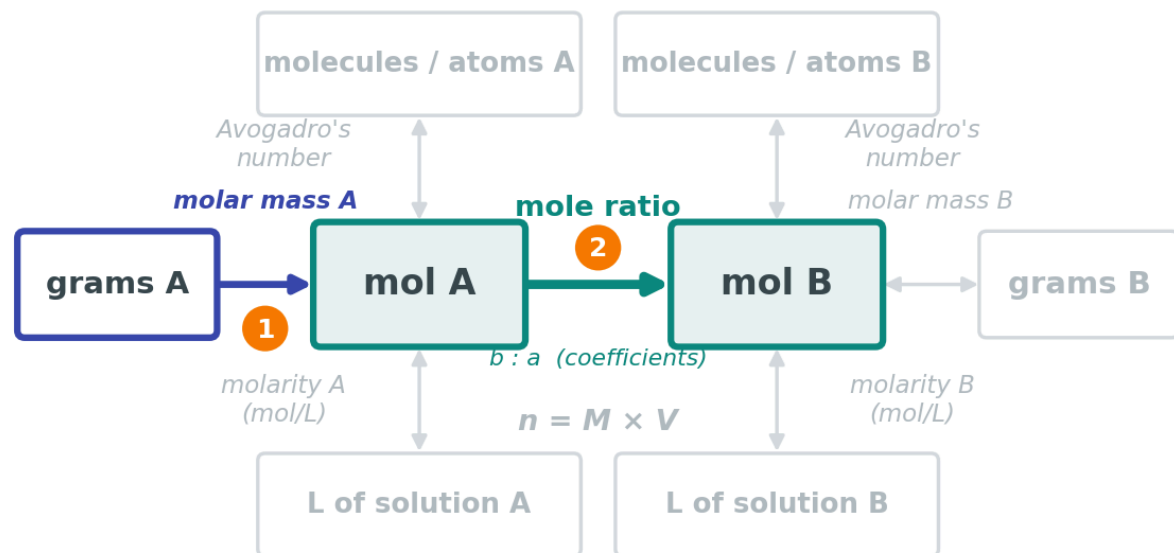
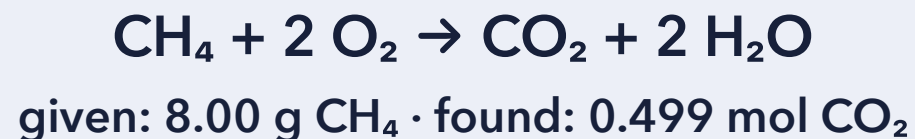
$$8.00 \text{ g } \cancel{\text{CH}_4} \times \frac{1 \text{ mol } \cancel{\text{CH}_4}}{16.04 \text{ g } \cancel{\text{CH}_4}} = 0.499 \text{ mol CH}_4$$

Step 2 · Moles → moles

$$8.00 \text{ g } \cancel{\text{CH}_4} \times \frac{1 \cancel{\text{ mol CH}_4}}{16.04 \text{ g } \cancel{\text{CH}_4}} \times \frac{1 \text{ mol CO}_2}{1 \cancel{\text{ mol CH}_4}} = \mathbf{0.499 \text{ mol CO}_2}$$

 **Sense check:** A 1 : 1 ratio carries the count across unchanged: 0.499 mol CH₄ makes 0.499 mol CO₂. The question asked for moles, and the chain stops here. ✓

✓ Worked example 1: the route on the map



Sense check: Two arrows: CH₄'s molar mass, then the mole ratio. The route stops at mol B because the question asked for moles. ✓

✓ Worked example 1: one more factor



found: 0.499 mol CO₂ from 8.00 g CH₄ · wanted: g CO₂ (44.01 g/mol)

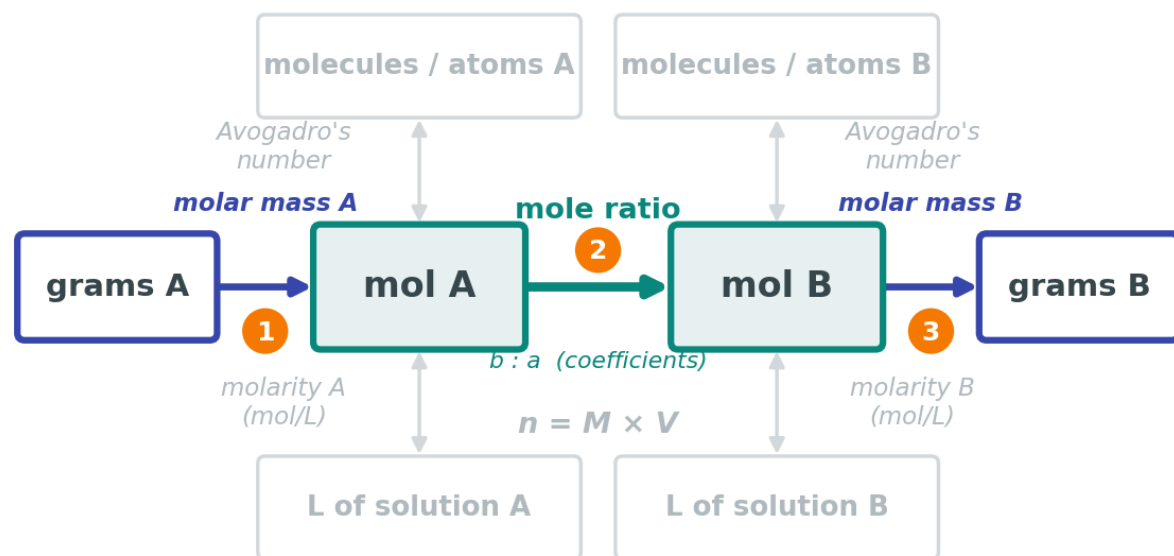
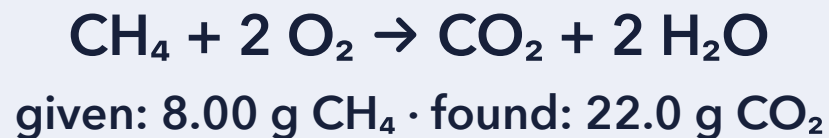
Step 3 · Moles → grams

CO₂'s molar mass converts the moles out to mass. The same chain, one factor longer:

$$8.00 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CH}_4} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 22.0 \text{ g CO}_2$$

 **Sense check:** Grams to grams is not a new method: it is the moles answer with a molar mass on the end. A mole of CO₂ (44.01 g) outweighs a mole of CH₄ (16.04 g), so 8.00 g of fuel becomes 22.0 g of CO₂. ✓

✓ Worked example 1: the grams route on the map



🧠 **Sense check:** Three arrows: the moles route plus CO₂'s molar mass at the end. ✓



Worked example 2



given: 90.0 g glucose · wanted: g CO₂

Respiration burns glucose, and the CO₂ is exhaled. What mass of CO₂ forms when **90.0 g of glucose** reacts completely? (*glucose 180.16 g/mol, CO₂ 44.01 g/mol*)

Write the route first: g glucose → mol glucose → mol CO₂ → g CO₂.

✓ Worked example 2: solution



given: 90.0 g glucose

Step 1 • Grams → moles

$$90.0 \text{ g glucose} \times \frac{1 \text{ mol glucose}}{180.16 \text{ g glucose}}$$

Step 2 • Moles → moles

Step 3 • Moles → grams

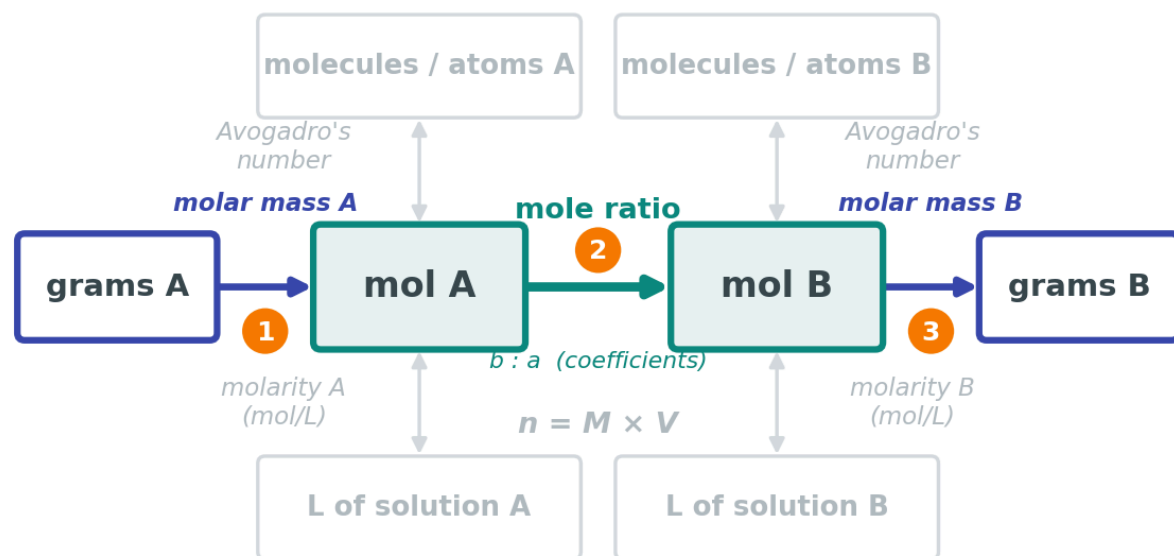
$$90.0 \text{ g glucose} \times \frac{1 \text{ mol glucose}}{180.16 \text{ g glucose}} \times \frac{6 \text{ mol CO}_2}{1 \text{ mol glucose}} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 132 \text{ g CO}_2$$

 **Sense check:** More mass leaves than entered: the carbon leaves as CO₂, and the added oxygen mass comes from the air. ✓

✓ Worked example 2: the route on the map



given: 90.0 g glucose · found: 132 g CO₂



🧠 **Sense check:** Three arrows, three conversion factors: glucose's molar mass, the mole ratio, CO₂'s molar mass. ✓

Your turn: potassium chlorate



Heating potassium chlorate releases oxygen gas. Starting from **61.3 g KClO₃** (122.55 g/mol; O₂ 32.00 g/mol):

$$61.3 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \times \frac{\text{ } \text{mol O}_2}{\text{ } \text{mol KClO}_3} \times \frac{\text{ } \text{g O}_2}{1 \text{ mol O}_2} = \text{ } \text{g O}_2$$

Fill the mole ratio from the coefficients and the last molar mass, then compute.

$$61.3 \text{ g KClO}_3 \times \frac{1 \text{ mol KClO}_3}{122.55 \text{ g KClO}_3} \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = \mathbf{24.0 \text{ g O}_2}$$

Where this goes wrong



given: 61.3 g KClO_3

Applying the mole ratio to grams. $61.3 \text{ g} \times (3/2) = 92.0 \text{ g}$ is wrong: coefficients count moles, not grams. Convert first: $61.3 \text{ g} \div 122.55 \text{ g/mol} = 0.500 \text{ mol}$. The correct answer is **24.0 g**.

Skipping the mole ratio. $\text{g} \rightarrow \text{mol} \rightarrow \text{g}$ gives 16.0 g. That assumes a 1:1 ratio. The equation gives 2 KClO_3 : 3 O_2 , and the mole ratio is the only step that switches substances.

Inverting the ratio. $(2 \text{ mol KClO}_3 / 3 \text{ mol O}_2)$ leaves units of $\text{mol KClO}_3^2/\text{mol O}_2$. Nothing cancels, and 10.7 g is wrong. If the units do not cancel, the fraction is inverted.



Practice 1



A camping stove burns **25.0 g of propane** (44.09 g/mol). What mass of water, in grams, forms? (H_2O 18.02 g/mol)

- A. 10.2
- B. 2.27
- C. 40.9
- D. 2.55

✓ Practice 1: answer C



given: 25.0 g C₃H₈

$$25.0 \text{ g C}_3\text{H}_8 \times \frac{1 \text{ mol C}_3\text{H}_8}{44.09 \text{ g C}_3\text{H}_8} \times \frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol C}_3\text{H}_8} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 40.9 \text{ g H}_2\text{O (answer C)}$$

A skipped the mole ratio: $25.0/44.09 \times 18.02 = 10.2$. B stopped at moles: $25.0/44.09 \times 4 = 2.27 \text{ mol H}_2\text{O}$, one factor short of grams. D inverted the ratio: $25.0/44.09 \times (1/4) \times 18.02 = 2.55$.

 **Sense check:** The ratio gives four H₂O per C₃H₈, but a mole of H₂O weighs less than half a mole of C₃H₈ (18 vs 44 g). The overall factor is about 1.6: 25.0 → 40.9 ✓



Practice 2



35.0 g of aluminum (26.98 g/mol) oxidizes completely. How many grams of Al_2O_3 (101.96 g/mol) form?

- A. 0.649
- B. 66.1
- C. 132
- D. 265

✓ Practice 2: answer B



given: 35.0 g Al

$$35.0 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{2 \text{ mol Al}_2\text{O}_3}{4 \text{ mol Al}} \times \frac{101.96 \text{ g Al}_2\text{O}_3}{1 \text{ mol Al}_2\text{O}_3} = 66.1 \text{ g Al}_2\text{O}_3 \text{ (answer B)}$$

A stopped at moles: $35.0/26.98 \times (2/4) = 0.649 \text{ mol Al}_2\text{O}_3$, one factor short of grams. C skipped the mole ratio: $35.0/26.98 \times 101.96 = 132$. D inverted the ratio: $35.0/26.98 \times (4/2) \times 101.96 = 265$.

 **Sense check:** The oxide weighs more than the metal alone. The extra 31.1 g is oxygen from the air. ✓



Practice 3



Baking soda releases CO_2 as it decomposes in a hot oven. A recipe needs **2.20 g of CO_2** to rise. How many grams of NaHCO_3 (84.01 g/mol) must decompose? (CO_2 44.01 g/mol)

- A. 4.20
- B. 2.10
- C. 0.100
- D. 8.40

✓ Practice 3: answer D



given: 2.20 g CO₂ · wanted: g NaHCO₃

$$2.20 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{2 \text{ mol NaHCO}_3}{1 \text{ mol CO}_2} \times \frac{84.01 \text{ g NaHCO}_3}{1 \text{ mol NaHCO}_3} = 8.40 \text{ g NaHCO}_3 \text{ (answer D)}$$

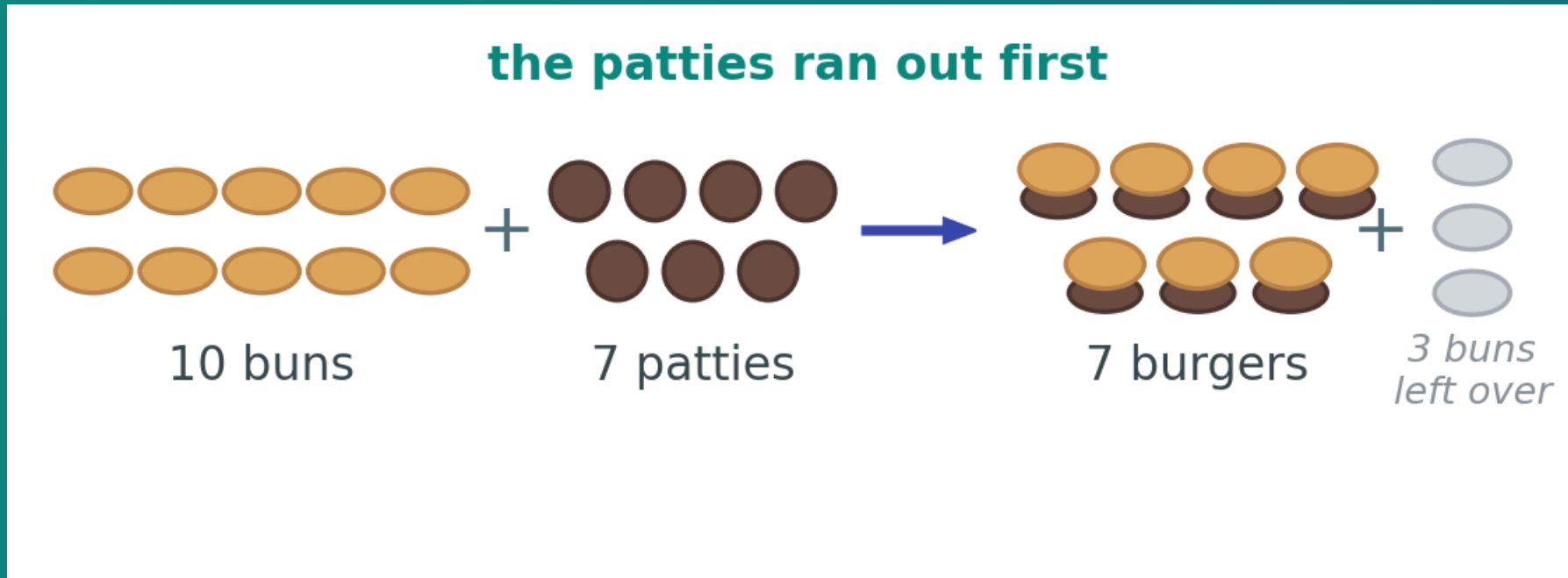
A skipped the mole ratio: $2.20/44.01 \times 84.01 = 4.20$. B inverted the ratio: $2.20/44.01 \times (1/2) \times 84.01 = 2.10$. C stopped at moles: $2.20/44.01 \times 2 = 0.100 \text{ mol NaHCO}_3$, one factor short of grams.

 **Sense check:** The route runs backward just as well: g CO₂ → mol CO₂ → mol NaHCO₃ → g NaHCO₃. Two NaHCO₃ per CO₂, each nearly twice as heavy: about four times the mass, $2.20 \rightarrow 8.40$ ✓

12 · Limiting Reactant

Decide which reactant runs out first, and compute the product from that reactant alone.

One ingredient runs out first

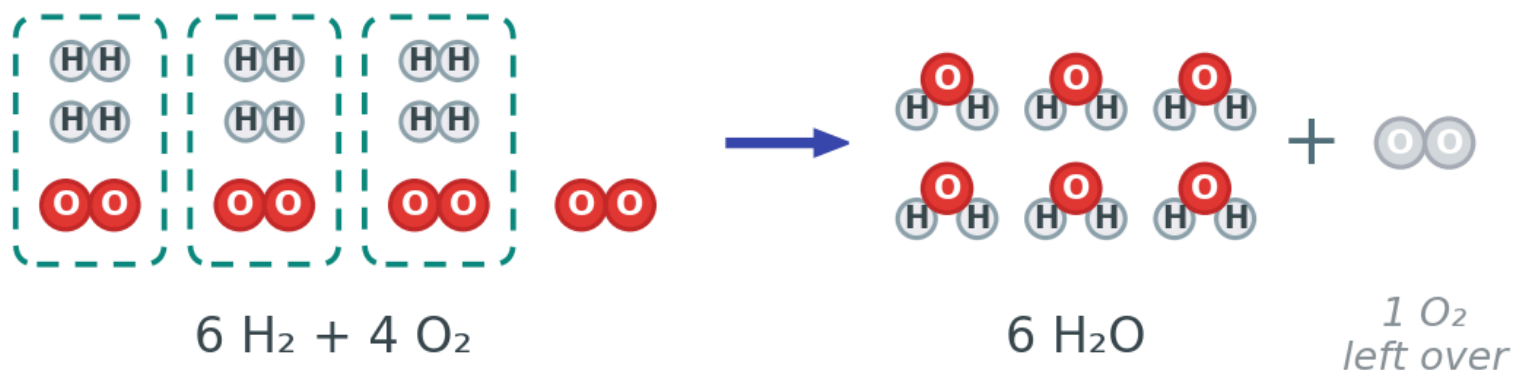


One bun and one patty per burger. The patties ran out first, so only seven burgers can be made. Reactions work the same way, in moles.

Limiting and excess



2 H₂ consumed per 1 O₂: the H₂ runs out first



Reactants are consumed in a **fixed ratio**, set by the coefficients. The first to run out is the **limiting reactant**: it stops the reaction and sets the maximum product. What remains of the other is **excess**.

Count it out



start: 6 H_2 and 4 O_2 · each run of the recipe consumes 2 H_2 and 1 O_2 , making 2 H_2O

Count the runs: how many H_2O form, and what remains?



run 1: 2 H_2 + 1 O_2 · run 2: 2 H_2 + 1 O_2 · run 3: 2 H_2 + 1 O_2 · no H_2 for a fourth
after: 6 H_2O made · 0 H_2 · 1 O_2 left over

Six H_2O form and one O_2 remains. H_2 started with more molecules and still ran out first.

The test: moles $\div$ coefficient



H_2 : $6.0 \div 2 = 3.0$ runs $\cdot$ O_2 : $4.0 \div 1 = 4.0$ runs $\cdot$ $3.0 < 4.0$, H_2 is spent first

Moles $\div$ coefficient counts how many times a reactant can run the recipe. The reactant with the fewest runs is spent first: the division predicts the count-out.

each run of the recipe consumes 2 mol H_2 + 1 mol O_2

	run 1	run 2	run 3	run 4	fewer runs: H_2 limits
H_2 6.0 mol	2 mol	2 mol	2 mol	empty	$6.0 \div 2 = 3.0$ runs
O_2 4.0 mol	1 mol	1 mol	1 mol	1 mol	$4.0 \div 1 = 4.0$ runs

no H_2 partner: excess O_2

the reaction stops here: H_2 is spent



The method

1. Convert **both** reactants to moles.
2. Divide each by its **coefficient**.
3. The **smaller** number limits: that reactant runs out first.
4. Compute product from the **limiting reactant only**.



Worked example 1: thermite

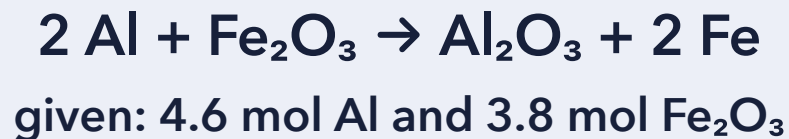


given: 4.6 mol Al and 3.8 mol Fe_2O_3

Rail-welding thermite: **4.6 mol Al** mixed with **3.8 mol Fe_2O_3** . How many moles of Al_2O_3 can form?

Divide each reactant's moles by its coefficient. The smaller result marks the limiting reactant.

✓ Worked example 1: which reactant limits



Step 1 · Convert both reactants to moles

Step 2 · Divide each by its coefficient

$$\text{Al: } \frac{4.6 \text{ mol}}{2} = 2.3$$

$$\text{Fe}_2\text{O}_3: \frac{3.8 \text{ mol}}{1} = 3.8$$

Step 3 · The smaller number limits

🧠 **Sense check:** $2.3 < 3.8$, so **Al limits**. Fe_2O_3 is excess.

✓ Worked example 1: how much product



given: 4.6 mol Al and 3.8 mol Fe_2O_3 · Al limits

Step 4 · Compute product from the limiting reactant only

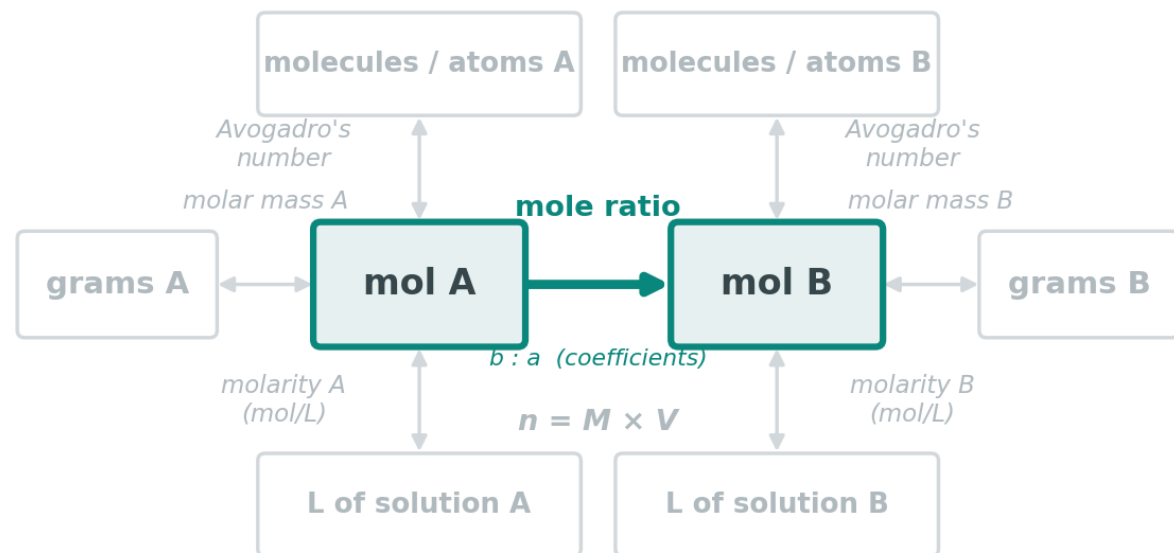
$$4.6 \text{ mol Al} \times \frac{1 \text{ mol Al}_2\text{O}_3}{2 \text{ mol Al}} = 2.3 \text{ mol Al}_2\text{O}_3$$

🧠 **Sense check:** Fe_2O_3 alone could give 3.8 mol, but the reaction stops when Al runs out, at 2.3 mol. The smaller result is the amount that can actually form. ✓

✓ Worked example 1: the route on the map



given: 4.6 mol Al and 3.8 mol Fe_2O_3 · found: 2.3 mol Al_2O_3



🧠 **Sense check:** Both amounts started in moles, so Step 1 used no arrow. Steps 2 and 3 happen off the map. In Step 4 only Al, the limiting reactant, crosses the bridge: A is Al, B is Al_2O_3 . ✓



Worked example 2: full mass chain



given: 42.0 g N₂ and 12.0 g H₂

An ammonia reactor is charged with **42.0 g N₂** and **12.0 g H₂**. What mass of NH₃ can form?

A common first attempt: H₂ has the smaller mass, 12.0 g, so H₂ limits. Test it.

✓ Worked example 2: which reactant limits



Step 1 • Convert both reactants to moles

$$42.0 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2} = 1.50 \text{ mol N}_2 \cdot 12.0 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} = 5.95 \text{ mol H}_2$$

Step 2 • Divide each by its coefficient

$$\text{N}_2: \frac{1.50 \text{ mol}}{1} = 1.50 \cdot \text{H}_2: \frac{5.95 \text{ mol}}{3} = 1.98$$

Step 3 • The smaller number limits

🧠 **Sense check:** $1.50 < 1.98$, so **N₂ limits**. H₂ is excess.

✓ Worked example 2: how much product



given: 42.0 g N₂ and 12.0 g H₂ · N₂ limits, 1.50 mol

Step 4 · Compute product from the limiting reactant only

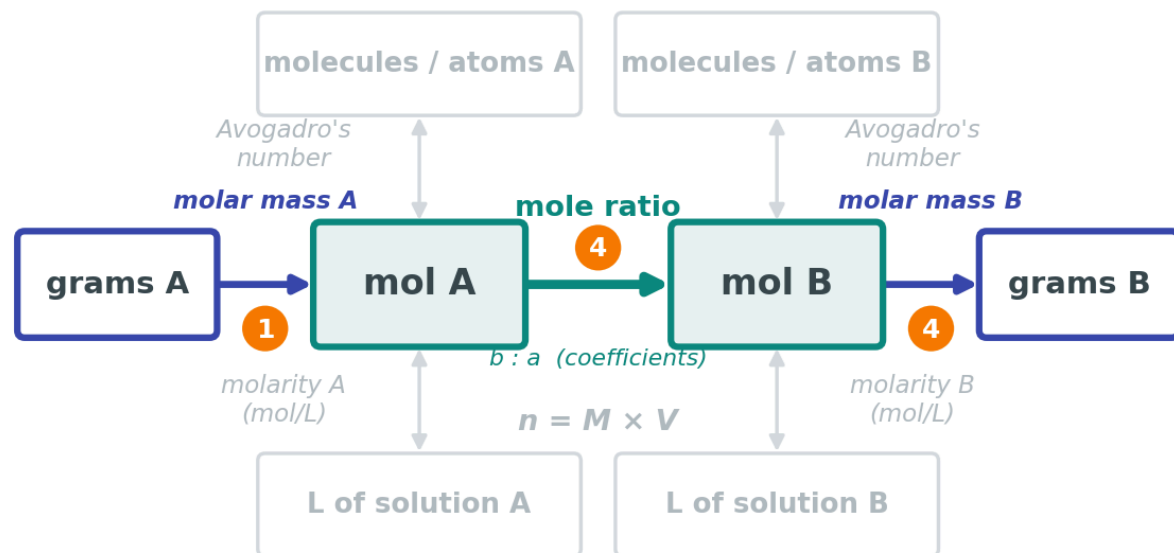
$$1.50 \text{ mol N}_2 \times \frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2} \times \frac{17.03 \text{ g NH}_3}{1 \text{ mol NH}_3} = 51.1 \text{ g NH}_3$$

 **Sense check:** The smaller mass, 12.0 g of H₂, is the excess; the larger mass, 42.0 g of N₂, runs out first. Mass does not identify the limiting reactant. ✓

✓ Worked example 2: the route on the map



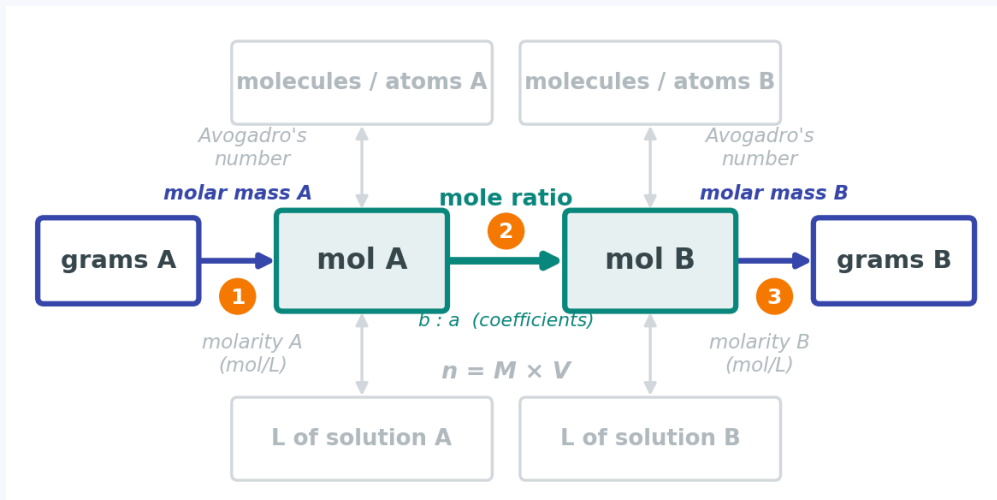
given: 42.0 g N₂ and 12.0 g H₂ · found: 51.1 g NH₃



🧠 **Sense check:** Step 1 runs arrow 1 twice (N₂, then H₂). Steps 2 and 3 happen off the map. In Step 4 only N₂, the limiting reactant, rides on to grams of NH₃. ✓



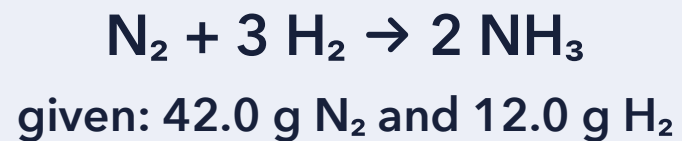
A second path: compare the product



1. Grams to moles, for **each** reactant.
2. Mole ratio to moles of product.
3. Moles to grams of product.
4. The reactant that makes **less** product limits. That amount is the **theoretical yield**.



Worked example 3: the product path



The same ammonia charge, solved a second way.

Convert each reactant to grams of NH₃ and compare the two results.

✓ Worked example 3: the NH_3 each reactant can make



given: 42.0 g N_2 and 12.0 g H_2

Step 1 • Grams to moles, for each reactant

$$42.0 \text{ g } \text{N}_2 \times \frac{1 \text{ mol } \text{N}_2}{28.02 \text{ g } \text{N}_2} = 1.50 \text{ mol } \text{N}_2 \cdot 12.0 \text{ g } \text{H}_2 \times \frac{1 \text{ mol } \text{H}_2}{2.016 \text{ g } \text{H}_2} = 5.95 \text{ mol } \text{H}_2$$

Step 2 • Mole ratio to moles of product

$$1.50 \text{ mol } \text{N}_2 \times \frac{2 \text{ mol } \text{NH}_3}{1 \text{ mol } \text{N}_2} = 3.00 \text{ mol } \text{NH}_3 \cdot 5.95 \text{ mol } \text{H}_2 \times \frac{2 \text{ mol } \text{NH}_3}{3 \text{ mol } \text{H}_2} = 3.97 \text{ mol } \text{NH}_3$$

 **Sense check:** Each reactant now has its own amount of NH_3 : 3.00 mol from N_2 , 3.97 mol from H_2 .

✓ Worked example 3: grams of NH₃, then compare



given: 42.0 g N₂ and 12.0 g H₂ · N₂ → 3.00 mol NH₃ · H₂ → 3.97 mol NH₃

Step 3 • Moles to grams of product

$$3.00 \text{ mol NH}_3 \times \frac{17.03 \text{ g NH}_3}{1 \text{ mol NH}_3} = 51.1 \text{ g NH}_3 \text{ from N}_2$$

$$3.97 \text{ mol NH}_3 \times \frac{17.03 \text{ g NH}_3}{1 \text{ mol NH}_3} = 67.6 \text{ g NH}_3 \text{ from H}_2$$

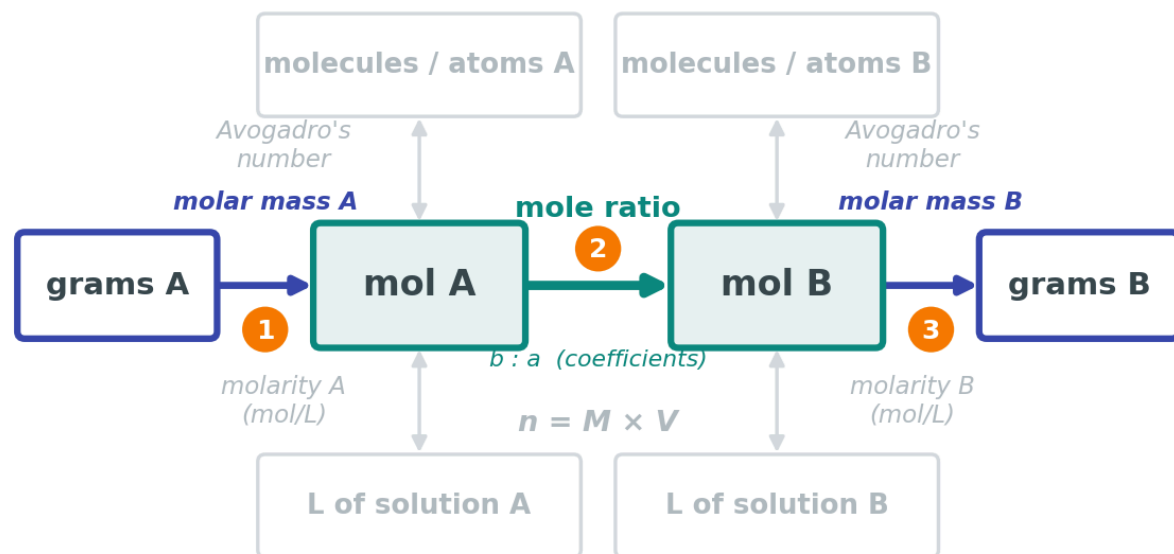
Step 4 • The reactant that makes less product limits

🧠 **Sense check:** 51.1 g < 67.6 g, so **N₂ limits**. The theoretical yield is **51.1 g NH₃**; the 67.6 g from H₂ can never form ✓

✓ Worked example 3: the route on the map



given: 42.0 g N₂ and 12.0 g H₂ · found: 51.1 g NH₃ from N₂, 67.6 g from H₂



 **Sense check:** The same three arrows, run once per reactant: A = N₂, then A = H₂; B = NH₃. The smaller grams of B names the limiting reactant. ✓

✓ Worked example 3: both paths agree



given: 42.0 g N₂ → 1.50 mol · 12.0 g H₂ → 5.95 mol

	moles ÷ coefficient	product path
N ₂	1.50 ÷ 1 = 1.50	1.50 × 2/1 = 3.00 mol NH ₃ → 51.1 g
H ₂	5.95 ÷ 3 = 1.98	5.95 × 2/3 = 3.97 mol NH ₃ → 67.6 g
verdict	N₂ limits	N₂ limits , 51.1 g NH ₃

 **Sense check:** Each product amount is the division result × 2, the coefficient of NH₃ (N₂: 1.50 × 2 = 3.00 mol). Multiplying both by the same number keeps the smaller one smaller, so both paths always name the same limiting reactant ✓

Your turn: finish the comparison



A signal flare holds **0.80 mol Mg** and **0.50 mol O₂**. How many moles of MgO form?

$$\text{Mg: } \frac{0.80 \text{ mol}}{2} = 0.40 \cdot \text{O}_2: \frac{0.50 \text{ mol}}{1} = \underline{\hspace{2cm}}$$

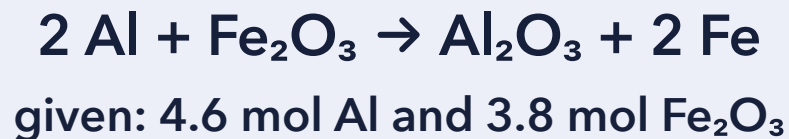
Limiting reactant:

$$\begin{aligned} \text{O}_2: 0.50 / 1 = \mathbf{0.50} \cdot \text{Mg: } 0.40 \text{ is smaller, so } \mathbf{\text{Mg limits}} \cdot 0.80 \cancel{\text{ mol Mg}} \times \frac{2 \text{ mol MgO}}{2 \cancel{\text{ mol Mg}}} \\ = \mathbf{0.80 \text{ mol MgO}} \end{aligned}$$

⚠ Where this goes wrong

Picking the limiting reactant by mass. In $\text{N}_2 + 3 \text{H}_2 \rightarrow 2 \text{NH}_3$, the smaller mass, 12.0 g of H_2 against 42.0 g of N_2 , is the excess. Mass does not identify the limiting reactant; moles $\div$ coefficient does.

Comparing raw moles. In $2 \text{Mg} + \text{O}_2 \rightarrow 2 \text{MgO}$, 0.50 mol O_2 is fewer moles than 0.80 mol Mg . But two Mg are consumed per O_2 : 0.40 vs 0.50, so **Mg** limits and **0.80 mol MgO** forms.



Building product from the excess reactant. With 4.6 mol Al and 3.8 mol Fe_2O_3 , the 3.8 mol suggests 3.8 mol Al_2O_3 . The reaction stops at 2.3 mol, when Al runs out. Compute product from the limiting reactant only.

Applying the mole ratio to grams. Coefficients count moles, not grams. Convert each mass to moles before comparing.

 Practice 1

10.0 g of Al foil reacts with **30.0 g of Cl₂** gas. What is the maximum mass of AlCl₃, in grams? (Al 26.98 g/mol · Cl₂ 70.90 g/mol · AlCl₃ 133.33 g/mol)

- A. 37.6
- B. 0.282
- C. 56.4
- D. 84.6

✓ Practice 1: answer A



given: 10.0 g Al, 30.0 g Cl₂ · Al: $0.371 \div 2 = 0.185$, excess · Cl₂: $0.423 \div 3 = 0.141 \leftarrow$ limits

$$30.0 \text{ g Cl}_2 \times \frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2} \times \frac{2 \text{ mol AlCl}_3}{3 \text{ mol Cl}_2} \times \frac{133.33 \text{ g AlCl}_3}{1 \text{ mol AlCl}_3} = 37.6 \text{ g AlCl}_3 \text{ (answer A)}$$

B stopped at moles: $30.0/70.90 \times (2/3) = 0.282 \text{ mol AlCl}_3$, one factor short of grams. C skipped the mole ratio: $30.0/70.90 \times 133.33 = 56.4$. D inverted the ratio: $30.0/70.90 \times (3/2) \times 133.33 = 84.6$.

 **Sense check:** Cl₂ brings three times the grams of Al, yet it runs out first: three Cl₂ at 70.90 g each are consumed per two Al ✓

 Practice 2

A methanol reactor is fed **14.0 g of CO** and **3.00 g of H₂**. What is the maximum mass of CH₃OH, in grams? (CO 28.01 g/mol · H₂ 2.016 g/mol · CH₃OH 32.04 g/mol)

- A. 23.8
- B. 16.0
- C. 449
- D. 0.500

✓ Practice 2: answer B



given: 14.0 g CO and 3.00 g H₂ · CO: 0.500 ÷ 1 = 0.500 ← limits · H₂: 1.49 ÷ 2 = 0.744, excess

$$14.0 \text{ g CO} \times \frac{1 \text{ mol CO}}{28.01 \text{ g CO}} \times \frac{1 \text{ mol CH}_3\text{OH}}{1 \text{ mol CO}} \times \frac{32.04 \text{ g CH}_3\text{OH}}{1 \text{ mol CH}_3\text{OH}} = 16.0 \text{ g CH}_3\text{OH (answer B)}$$

A took the smaller mass, H₂, as limiting: 3.00/2.016 × (1/2) × 32.04 = 23.8. C skipped the grams-to-moles step and treated grams of CO as moles: 14.0 × 32.04 = 449. D stopped at moles: 14.0/28.01 × (1/1) = 0.500 mol CH₃OH, one factor short of grams.

🧠 **Sense check:** Mass balance: 0.98 g H₂ is left over, and 16.0 + 0.98 = 17.0 g loaded ✓

Practice 3: the product path



A CO₂ scrubber cartridge holds **36.0 g of NaOH** and takes in **22.0 g of CO₂**. What is the theoretical yield of Na₂CO₃, in grams? (*NaOH 40.00 g/mol · CO₂ 44.01 g/mol · Na₂CO₃ 105.99 g/mol*)

- A. 53.0
- B. 0.450
- C. 47.7
- D. 191
- E. 95.4

✓ Practice 3: answer C



given: 36.0 g NaOH $\rightarrow$ 0.900 mol $\cdot$ 22.0 g CO₂ $\rightarrow$ 0.500 mol

$$0.900 \text{ mol NaOH} \times \frac{1 \text{ mol Na}_2\text{CO}_3}{2 \text{ mol NaOH}} \times \frac{105.99 \text{ g Na}_2\text{CO}_3}{1 \text{ mol Na}_2\text{CO}_3} = 47.7 \text{ g Na}_2\text{CO}_3 \text{ from NaOH}$$

$$0.500 \text{ mol CO}_2 \times \frac{1 \text{ mol Na}_2\text{CO}_3}{1 \text{ mol CO}_2} \times \frac{105.99 \text{ g Na}_2\text{CO}_3}{1 \text{ mol Na}_2\text{CO}_3} = 53.0 \text{ g Na}_2\text{CO}_3 \text{ from CO}_2$$

🧠 **Sense check:** 47.7 g < 53.0 g, so NaOH limits: theoretical yield **47.7 g Na₂CO₃ (answer C)** ✓

A kept the larger amount, from CO₂, the excess (also the smaller mass and fewer moles): 53.0. B stopped at moles: 0.900 $\times$ (1/2) = 0.450. D inverted the ratio: 0.900 $\times$ 2 $\times$ 105.99 = 191. E skipped the ratio: 0.900 $\times$ 105.99 = 95.4.

 Practice 5

A titanium reactor is charged with **50.00 g of TiCl_4** and **20.00 g of Mg** . What mass of the excess reactant, in grams, is left over when the reaction stops? (TiCl_4 189.67 g/mol · Mg 24.31 g/mol)

- A. 12.82
- B. 16.80
- C. 13.59
- D. 7.18

✓ Practice 5: answer D



given: 50.00 g TiCl_4 and 20.00 g Mg · TiCl_4 : $0.2636 \div 1 = 0.2636 \leftarrow$ limits · Mg: $0.8227 \div 2 = 0.4114$,
excess

$$50.00 \text{ g } \text{TiCl}_4 \times \frac{1 \text{ mol } \text{TiCl}_4}{189.67 \text{ g } \text{TiCl}_4} \times \frac{2 \text{ mol Mg}}{1 \text{ mol } \text{TiCl}_4} \times \frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = 12.82 \text{ g Mg consumed}$$

$$20.00 \text{ g loaded} - 12.82 \text{ g consumed} = \mathbf{7.18 \text{ g Mg left over (answer D)}}$$

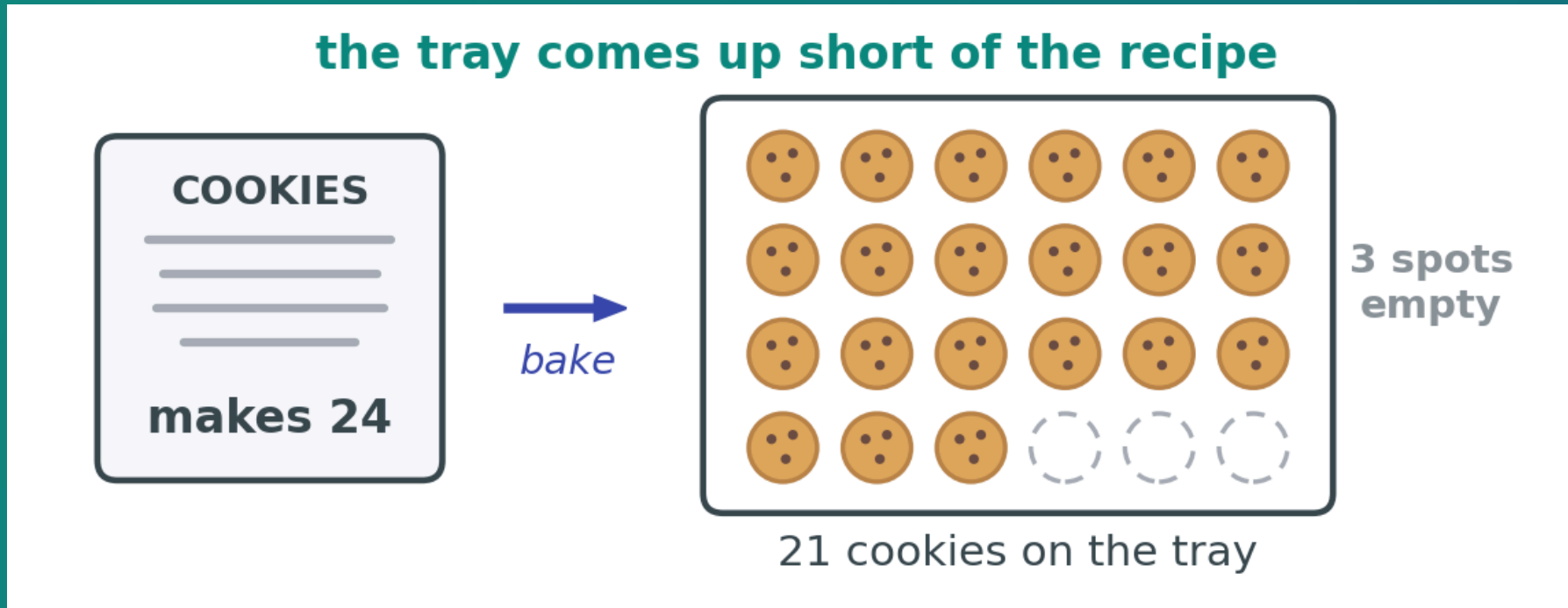
A stopped at the Mg consumed, 12.82 g. B inverted the ratio: $20.00 - 50.00/189.67 \times (1/2) \times 24.31 = 16.80$. C skipped it: $20.00 - 50.00/189.67 \times 24.31 = 13.59$.

🧠 **Sense check:** The smaller mass, Mg, is the excess. The leftover lies between zero and the 20.00 g loaded ✓

13 · Percent Yield

Compare the mass a reaction actually delivers to the maximum stoichiometry allows, and report it as a percent yield.

Three cookies short



A cookie recipe promises two dozen. The tray comes out with 21: batter stuck to the bowl, one burned. Chemical reactions come up short the same way.

Theoretical yield: the maximum stoichiometry allows



given: 10.0 g Mg · the chain g Mg → mol Mg → mol MgO → g MgO allows at most 16.6 g

Mass-to-mass stoichiometry computes the most product the given amounts can form: the **theoretical yield**. A real experiment collects less. **Percent yield** reports how much of that maximum was actually delivered.

Actual yield is measured on the balance

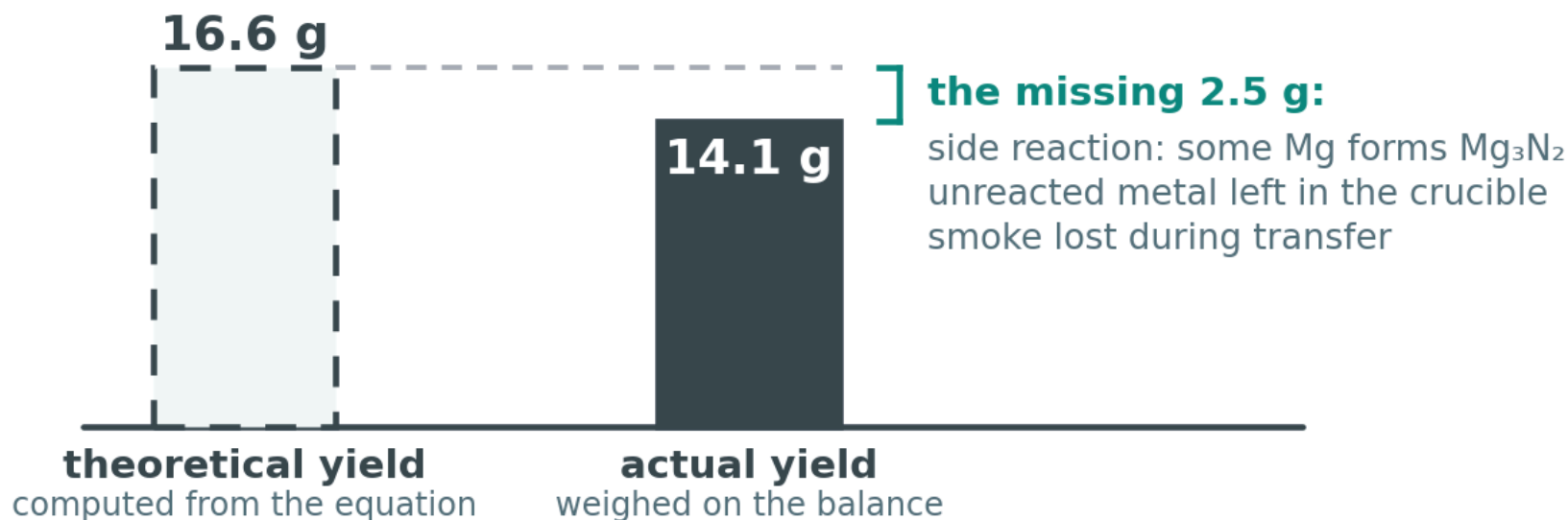


theoretical: 16.6 g MgO (computed) · actual: 14.1 g MgO (weighed)

The **actual yield** is the mass of product collected, read off the balance after the experiment. A problem either states it or gives a typical percent yield that predicts it.

Where the missing mass goes

the balance reads less than the maximum



Side reactions consume reactant without making the product. Some reactant never reacts. Some product stays behind in transfer: on the filter, in the crucible. Every loss lowers the actual yield.

Percent yield

$$\text{percent yield} = \text{actual yield} \div \text{theoretical yield} \times 100$$

actual = weighed · theoretical = computed · same substance, same unit

$$\text{actual yield} = \text{percent yield} \div 100 \times \text{theoretical yield}$$

the same equation solved for actual · a known typical yield predicts what a run delivers

Both masses refer to the product. The fraction compares the collected mass to the maximum; $\times 100$ states it as a percent. A real preparation lands below 100%.



The method

1. **Compute the theoretical yield**: the maximum product mass. Two reactant amounts given: work from the limiting reactant.
2. **Take the actual yield from the problem**: stated, or the unknown.
3. **Divide**: $\text{actual} \div \text{theoretical} \times 100$.

$$\text{actual yield} = \text{percent yield} \div 100 \times \text{theoretical yield}$$

percent given, actual wanted: step 3 solved for actual



Worked example 1: heating limestone



given: 50.0 g CaCO_3 · actual: 18.5 g CO_2 collected · wanted: percent yield

A kiln charge of **50.0 g CaCO_3** is heated until no more gas comes off. The CO_2 collected weighs **18.5 g**. What is the percent yield? (CaCO_3 100.09 g/mol · CO_2 44.01 g/mol)

A common first attempt: $18.5 \div 50.0 \times 100 = 37.0\%$. Test it.

✓ Worked example 1: solution



given: 50.0 g CaCO_3 · actual: 18.5 g CO_2

Step 1 · Compute the theoretical yield

$$50.0 \text{ g } \text{CaCO}_3 \times \frac{1 \text{ mol } \text{CaCO}_3}{100.09 \text{ g } \text{CaCO}_3} \times \frac{1 \text{ mol } \text{CO}_2}{1 \text{ mol } \text{CaCO}_3} \times \frac{44.01 \text{ g } \text{CO}_2}{1 \text{ mol } \text{CO}_2} = 22.0 \text{ g } \text{CO}_2$$

Step 2 · Take the actual yield from the problem

Step 3 · Divide: actual ÷ theoretical × 100

$$\frac{18.5 \text{ g } \text{CO}_2}{22.0 \text{ g } \text{CO}_2} \times 100 = 84.1\%$$

 **Sense check:** Below 100% ✓. Of every 100 g of CO_2 the equation allows, the kiln delivered 84. The 37.0% first attempt compared product to reactant, not product to product.



Worked example 2: two reactant amounts



given: 15.0 g $\text{Pb}(\text{NO}_3)_2$ and 12.0 g KI · actual: 14.2 g PbI_2 · wanted: percent yield

Two clear solutions are mixed and bright yellow PbI_2 precipitates. The dried solid weighs **14.2 g**. What is the percent yield? ($\text{Pb}(\text{NO}_3)_2$ 331.2 g/mol · KI 166.00 g/mol · PbI_2 461.0 g/mol)

Amounts of both reactants are given. The theoretical yield comes from the limiting reactant.

✓ Worked example 2: which reactant limits



given: 15.0 g $\text{Pb}(\text{NO}_3)_2$ and 12.0 g KI · actual: 14.2 g PbI_2

Step 1 · Compute the theoretical yield

$$15.0 \text{ g } \cancel{\text{Pb}(\text{NO}_3)_2} \times \frac{1 \text{ mol } \text{Pb}(\text{NO}_3)_2}{331.2 \text{ g } \cancel{\text{Pb}(\text{NO}_3)_2}} = 0.0453 \text{ mol} \cdot 12.0 \text{ g } \cancel{\text{KI}} \times \frac{1 \text{ mol KI}}{166.00 \text{ g } \cancel{\text{KI}}} \\ = 0.0723 \text{ mol}$$

$$\text{Pb}(\text{NO}_3)_2: \frac{0.0453 \text{ mol}}{1} = 0.0453 \text{ (excess)} \cdot \text{KI}: \frac{0.0723 \text{ mol}}{2} = \mathbf{0.0362} \leftarrow \text{smaller: KI limits}$$

🧠 **Sense check:** $0.0362 < 0.0453$: KI runs out first. The theoretical yield comes from KI alone.

✓ Worked example 2: the percent yield



given: 15.0 g $\text{Pb}(\text{NO}_3)_2$ and 12.0 g KI · KI limits, 0.0723 mol · actual: 14.2 g PbI_2

Step 1 · Compute the theoretical yield

$$12.0 \text{ g KI} \times \frac{1 \text{ mol KI}}{166.00 \text{ g KI}} \times \frac{1 \text{ mol PbI}_2}{2 \text{ mol KI}} \times \frac{461.0 \text{ g PbI}_2}{1 \text{ mol PbI}_2} = 16.66 \text{ g PbI}_2 \text{ theoretical}$$

Step 2 · Take the actual yield from the problem

Step 3 · Divide: actual ÷ theoretical × 100

$$\frac{14.2 \text{ g PbI}_2}{16.66 \text{ g PbI}_2} \times 100 = 85.2\%$$

 **Sense check:** Below 100% ✓. Some PbI_2 stayed dissolved and some clung to the filter: the balance reads less than the maximum.

Your turn: blast furnace iron



given: 80.0 g Fe_2O_3 , CO in excess · actual: 47.6 g Fe

$$80.0 \text{ g Fe}_2\text{O}_3 \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.70 \text{ g Fe}_2\text{O}_3} \times \frac{\text{ } \text{mol Fe}}{\text{ } \text{mol Fe}_2\text{O}_3} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = \text{ } \text{g Fe}$$

$$\text{percent yield: } \frac{47.6 \text{ g Fe}}{\text{ } \text{g Fe}} \times 100 = \text{ }$$

$$\text{ratio: } \frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} \cdot \text{theoretical: } 55.95 \text{ g Fe} \cdot \frac{47.6 \text{ g Fe}}{55.95 \text{ g Fe}} \times 100 = 85.1\%$$

⚠ Where this goes wrong

Swapping actual and theoretical. With 14.2 g of PbI_2 collected and 16.66 g possible, $16.66 \div 14.2 \times 100 = 117\%$. No experiment beats its maximum; a percent above 100 means the ratio is upside down. Actual goes on top: $14.2 \div 16.66 \times 100 = 85.2\%$.

Computing the theoretical yield from the excess reactant. In $\text{Pb}(\text{NO}_3)_2 + 2 \text{KI} \rightarrow \text{PbI}_2 + 2 \text{KNO}_3$, with 15.0 g $\text{Pb}(\text{NO}_3)_2$ and 12.0 g KI, the $\text{Pb}(\text{NO}_3)_2$ chain gives 20.88 g and $14.2 \div 20.88 \times 100 = 68.0\%$. KI runs out first: the maximum is 16.66 g and the yield is 85.2%.

Comparing product to starting material. In $\text{Fe}_2\text{O}_3 + 3 \text{CO} \rightarrow 2 \text{Fe} + 3 \text{CO}_2$, 47.6 g of Fe from 80.0 g of Fe_2O_3 suggests $47.6 \div 80.0 \times 100 = 59.5\%$. The 100% mark is the 55.95 g of Fe stoichiometry allows, and the yield is 85.1%.

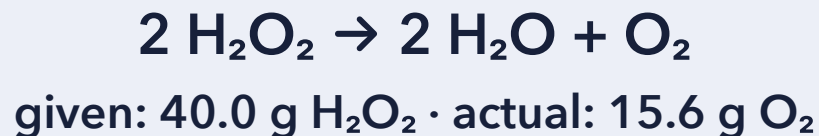
Reporting the fraction as the percent. $18.5 \div 22.0 = 0.841$ is a fraction of the maximum. Multiplied by 100 it becomes the percent yield, 84.1%. An answer of 0.841% would mean nearly everything was lost.

 Practice 1

A bottle of hydrogen peroxide decomposes completely. From **40.0 g of H_2O_2** , **15.6 g of O_2** is collected. What is the percent yield? (H_2O_2 34.02 g/mol · O_2 32.00 g/mol)

- A. 121
- B. 82.9
- C. 0.829
- D. 41.5

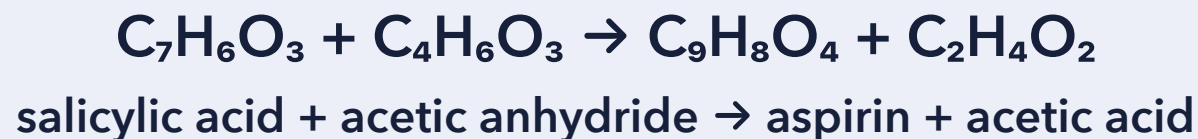
✓ Practice 1: answer B



$$40.0 \text{ g } \text{H}_2\text{O}_2 \times \frac{1 \text{ mol } \text{H}_2\text{O}_2}{34.02 \text{ g } \text{H}_2\text{O}_2} \times \frac{1 \text{ mol } \text{O}_2}{2 \text{ mol } \text{H}_2\text{O}_2} \times \frac{32.00 \text{ g } \text{O}_2}{1 \text{ mol } \text{O}_2} = 18.81 \text{ g } \text{O}_2 \text{ theoretical}$$

$$\frac{15.6 \text{ g } \text{O}_2}{18.81 \text{ g } \text{O}_2} \times 100 = \mathbf{82.9\% \text{ (answer B)}}$$

 **Sense check:** The product is a gas, and some escapes collection. About 83 g reached the flask for every 100 g possible. Below 100% ✓

 Practice 2

A student mixes **2.00 g of salicylic acid** with **5.00 g of acetic anhydride** and collects **2.15 g of aspirin**. What is the percent yield? ($\text{C}_7\text{H}_6\text{O}_3$ 138.12 g/mol $\cdot$ $\text{C}_4\text{H}_6\text{O}_3$ 102.09 g/mol $\cdot$ $\text{C}_9\text{H}_8\text{O}_4$ 180.16 g/mol)

- A. 24.4
- B. 121
- C. 0.824
- D. 82.4

✓ Practice 2: answer D



given: 2.00 g $\text{C}_7\text{H}_6\text{O}_3$ and 5.00 g $\text{C}_4\text{H}_6\text{O}_3$ · actual: 2.15 g $\text{C}_9\text{H}_8\text{O}_4$ · $\text{C}_7\text{H}_6\text{O}_3$: $0.01448 \div 1 \leftarrow \text{limits}$ ·
 $\text{C}_4\text{H}_6\text{O}_3$: $0.04898 \div 1$, excess

$$2.00 \text{ g } \text{C}_7\text{H}_6\text{O}_3 \times \frac{1 \text{ mol } \text{C}_7\text{H}_6\text{O}_3}{138.12 \text{ g } \text{C}_7\text{H}_6\text{O}_3} \times \frac{1 \text{ mol } \text{C}_9\text{H}_8\text{O}_4}{1 \text{ mol } \text{C}_7\text{H}_6\text{O}_3} \times \frac{180.16 \text{ g } \text{C}_9\text{H}_8\text{O}_4}{1 \text{ mol } \text{C}_9\text{H}_8\text{O}_4}$$

= 2.61 g theoretical

$$\frac{2.15 \text{ g } \text{C}_9\text{H}_8\text{O}_4}{2.61 \text{ g } \text{C}_9\text{H}_8\text{O}_4} \times 100 = \mathbf{82.4\% \text{ (answer D)}}$$

 **Sense check:** Below 100% ✓. Aspirin is recrystallized to purify it, and some stays dissolved in the rinse: the maximum comes from the limiting salicylic acid.

 Practice 3

A miner's lamp drips water onto **12.8 g of CaC_2** ; water is in excess. The reaction runs at **80.0%** yield. What mass of C_2H_2 , in grams, does the lamp actually produce? (CaC_2 64.10 g/mol · C_2H_2 26.04 g/mol)

- A. 4.16
- B. 5.20
- C. 6.50
- D. 10.2

✓ Practice 3: answer A



given: 12.8 g CaC_2 , water in excess · percent yield: 80.0% · wanted: actual g C_2H_2

$$12.8 \text{ g } \text{CaC}_2 \times \frac{1 \text{ mol } \text{CaC}_2}{64.10 \text{ g } \text{CaC}_2} \times \frac{1 \text{ mol } \text{C}_2\text{H}_2}{1 \text{ mol } \text{CaC}_2} \times \frac{26.04 \text{ g } \text{C}_2\text{H}_2}{1 \text{ mol } \text{C}_2\text{H}_2} = 5.20 \text{ g theoretical}$$

$$\text{actual} = \frac{80.0}{100} \times 5.20 \text{ g} = \mathbf{4.16 \text{ g } \text{C}_2\text{H}_2 \text{ (answer A)}}$$

🧠 **Sense check:** Percent yield = actual ÷ theoretical × 100, solved for actual. An 80.0% yield must land below the 5.20 g maximum: 4.16 g ✓

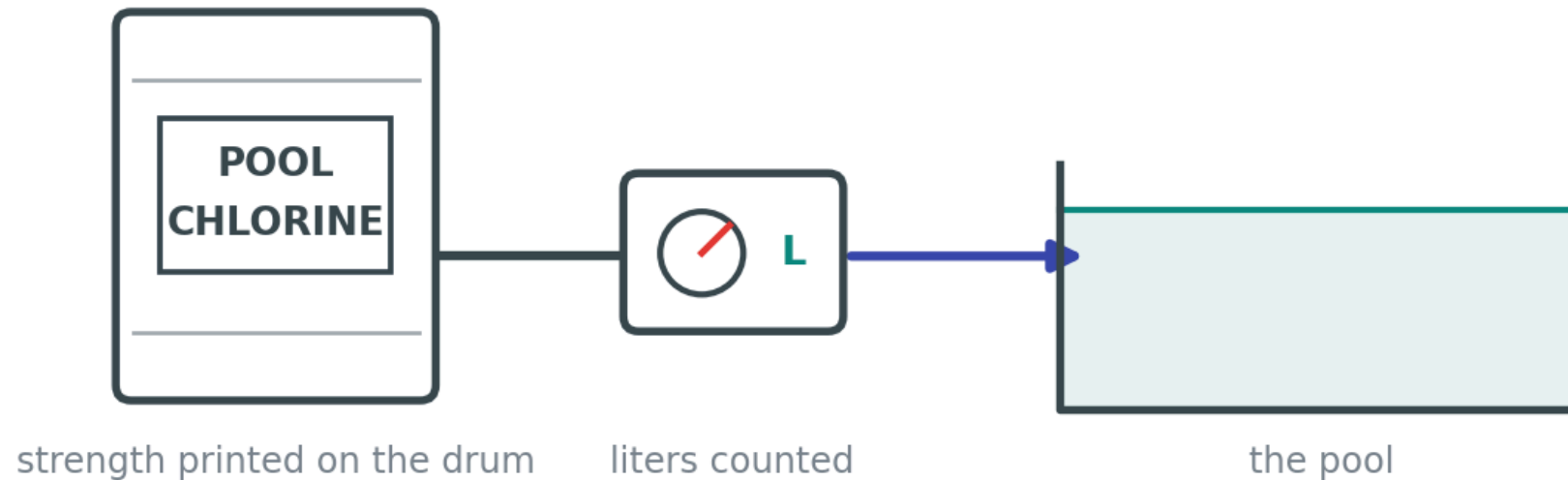
14 · Solution Stoichiometry

Convert a volume of solution to moles with molarity, cross substances with the mole ratio, and finish in grams or in the volume of a second solution.

Dosed by volume

dose = liters pumped × strength on the label

nothing is weighed



A pool is chlorinated by pumping in a measured volume of chlorine solution. Nothing is weighed: the strength on the drum's label and the liters pumped set the chlorine dose.

A solution's label counts moles

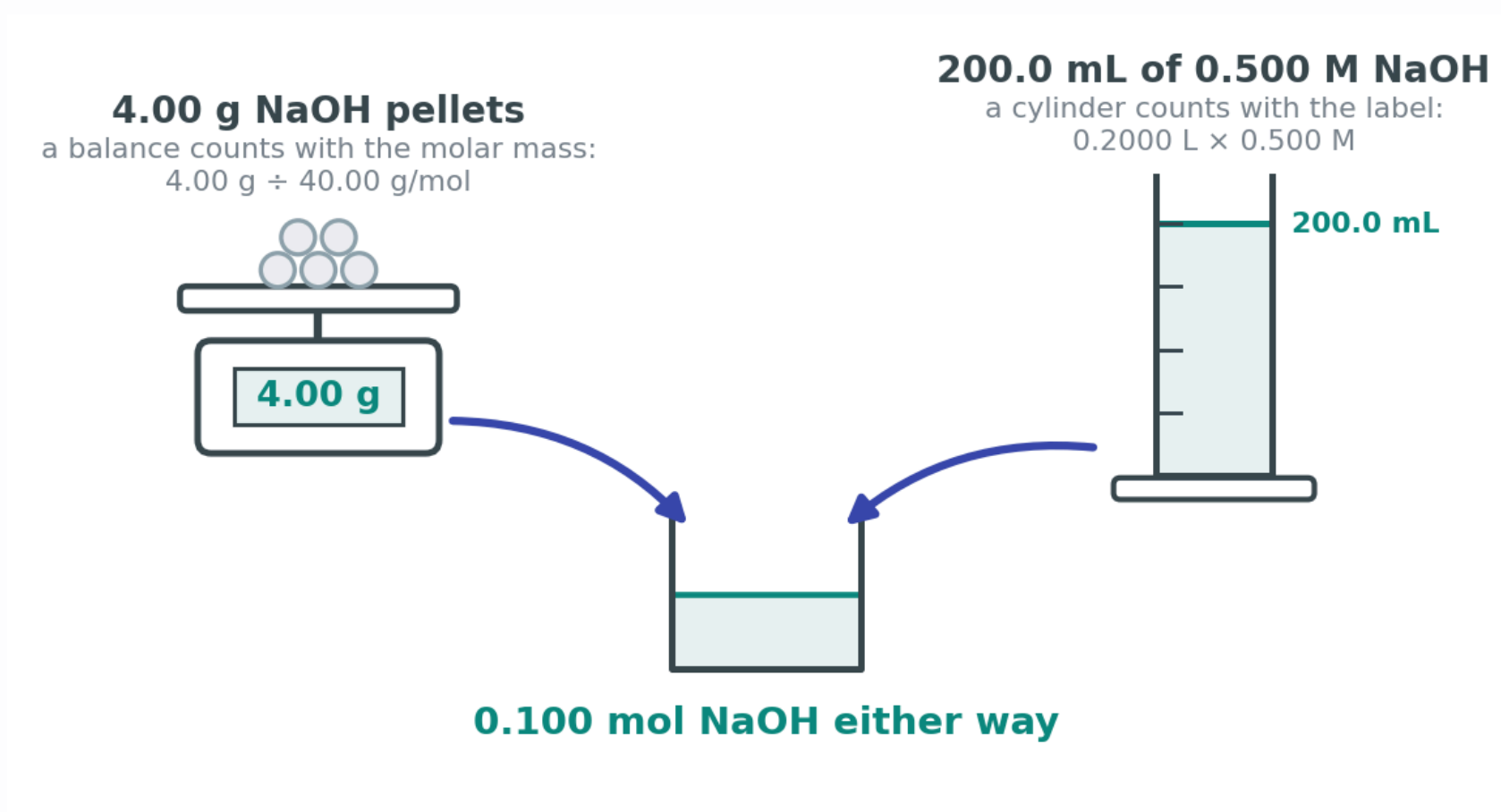
0.500 M NaOH: 0.500 mol NaOH = 1 L of solution
the label states a rate: moles delivered per liter poured

Measure a volume, and the label converts it to moles; the liters cancel:

$$0.2000 \text{ L soln} \times \frac{0.500 \text{ mol NaOH}}{1 \text{ L soln}} = \mathbf{0.100 \text{ mol NaOH}}$$

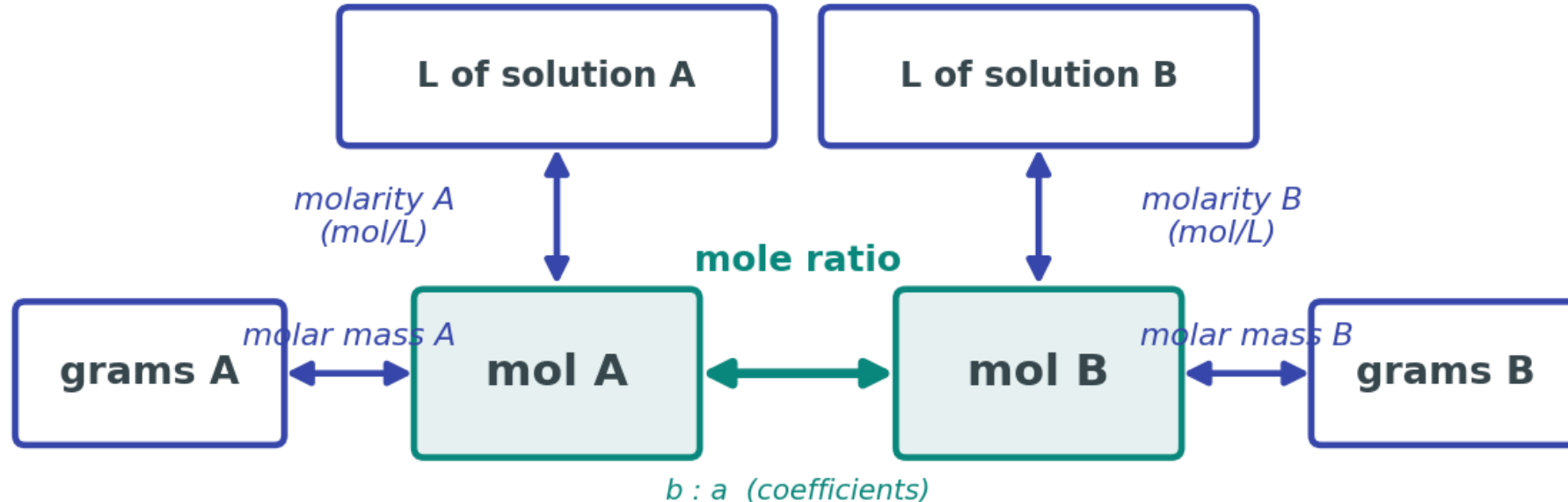
A graduated cylinder now counts moles. No balance is needed.

Two ways to deliver 0.100 mol



A reaction receives 0.100 mol either way. Stoichiometry works on moles; where they came from never enters the calculation.

Volume joins the map



a volume and a label count moles: no balance needed

Volume of solution enters through molarity, exactly where grams enter through molar mass. Every route still crosses the mole bridge, and the mole ratio still switches substances.



The method

1. **Volume → moles:** mL to L; molarity converts liters to moles of its solute.
2. **Moles → moles:** cross substances with the mole ratio. No other step can.
3. **Moles → the wanted unit:** molar mass for grams; molarity for volume.



Worked example 1: moles from a solution volume



given: 250.0 mL of 0.400 M HCl · wanted: mol H₂

Zinc metal dissolves in hydrochloric acid, releasing hydrogen gas. **250.0 mL of 0.400 M HCl** reacts completely with excess zinc. How many moles of H₂ form?

Write the route first: mL → L → mol HCl → mol H₂.

✓ Worked example 1: solution



given: 250.0 mL of 0.400 M HCl · wanted: mol H₂

Step 1 · Volume → moles

$$0.2500 \text{ L soln} \times \frac{0.400 \text{ mol HCl}}{1 \text{ L soln}} = 0.100 \text{ mol HCl}$$

Step 2 · Moles → moles

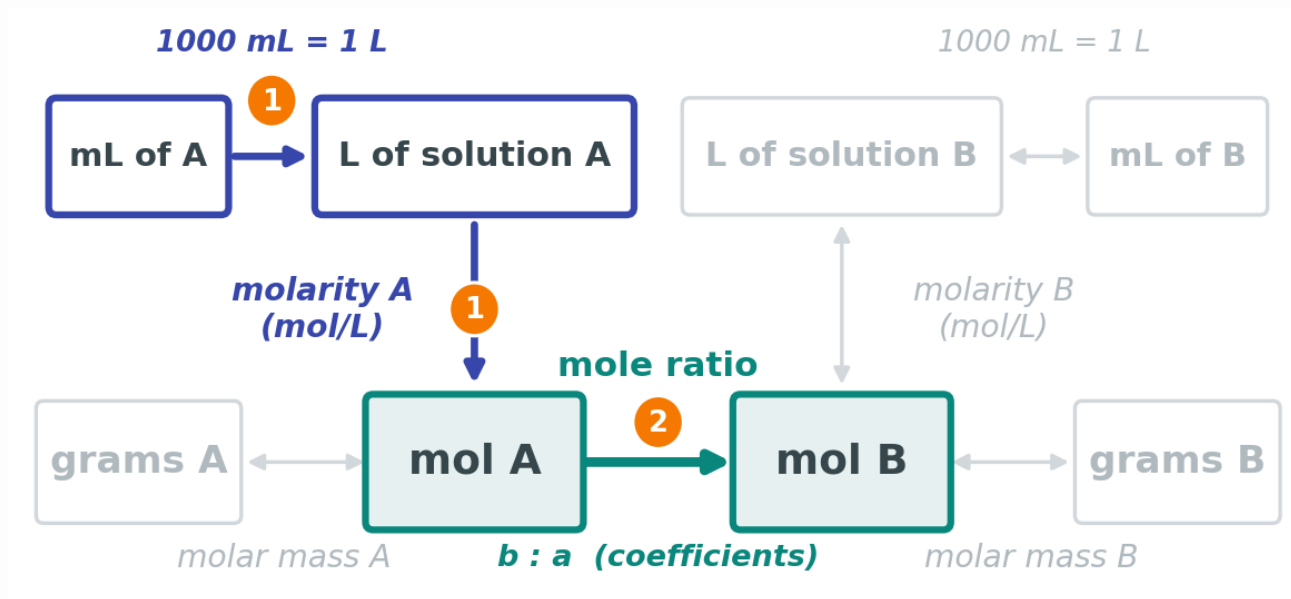
$$0.2500 \text{ L soln} \times \frac{0.400 \text{ mol HCl}}{1 \text{ L soln}} \times \frac{1 \text{ mol H}_2}{2 \text{ mol HCl}} = 0.0500 \text{ mol H}_2$$

 **Sense check:** Each H₂ consumes two HCl, so the mole count halves: 0.100 mol HCl → 0.0500 mol H₂ ✓

✓ Worked example 1: the route on the map



given: 250.0 mL of 0.400 M HCl · found: 0.0500 mol H₂



🧠 **Sense check:** A is HCl, B is H₂. Moles were wanted, so the route stops at mol B after the mole ratio. ✓

Worked example 2: grams of product from a solution volume



Two clear solutions mix, and bright yellow lead(II) iodide settles out. **150.0 mL of 0.300 M KI** reacts completely with excess Pb(NO₃)₂. What mass of PbI₂ forms? (*PbI₂ 461.0 g/mol*)

A tempting shortcut: carry the moles of KI straight to grams of PbI₂. Test it against the equation.

✓ Worked example 2: solution



given: 150.0 mL of 0.300 M KI · wanted: g PbI₂ (461.0 g/mol)

Step 1 · Volume → moles

The tempting shortcut

$$0.0450 \text{ mol KI} \times \frac{461.0 \text{ g PbI}_2}{1 \text{ mol PbI}_2} = 20.7 \text{ g} \times$$

Step 2 · Moles → moles

Step 3 · Moles → the wanted unit

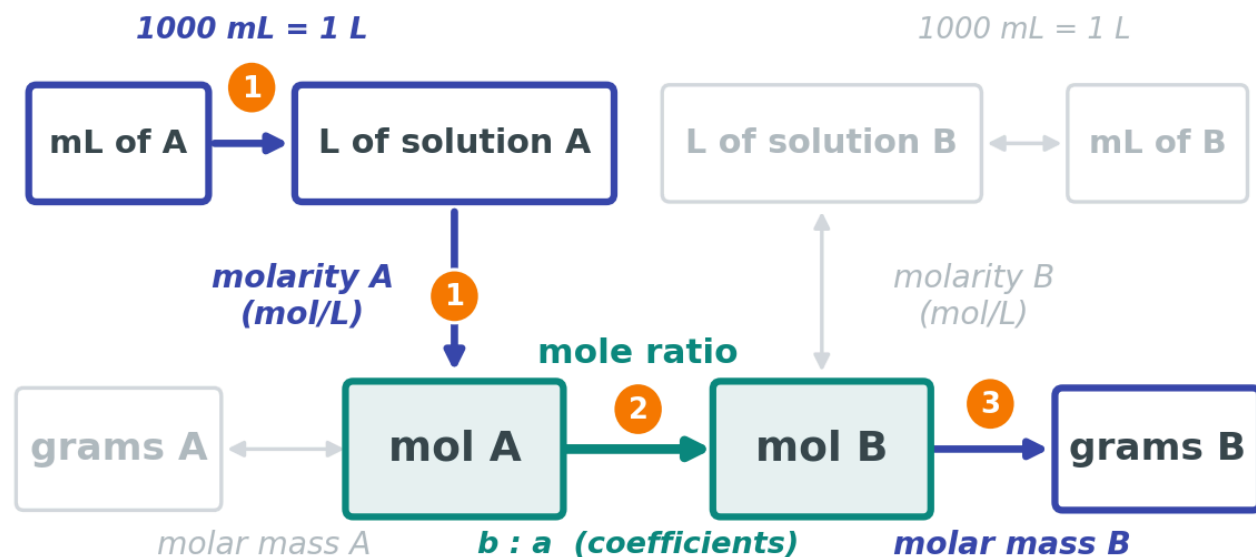
$$0.1500 \text{ L soln} \times \frac{0.300 \text{ mol KI}}{1 \text{ L soln}} \times \frac{1 \text{ mol PbI}_2}{2 \text{ mol KI}} \times \frac{461.0 \text{ g PbI}_2}{1 \text{ mol PbI}_2} = 10.4 \text{ g PbI}_2$$

 **Sense check:** Two KI deliver one PbI₂: 0.0450 mol halves to 0.0225 mol, and each mole weighs 461.0 g: about 10 g. 10.4 g ✓

✓ Worked example 2: the route on the map



given: 150.0 mL of 0.300 M KI · found: 10.4 g PbI_2



🧠 **Sense check:** A is KI, B is PbI_2 . The tempting shortcut jumped from mol A to grams B, but no arrow joins them: every route crosses the mole ratio. ✓

Your turn: copper(II) hydroxide



given: 200.0 mL (0.2000 L) of 0.250 M NaOH · wanted: g Cu(OH)₂ (97.57 g/mol)

$$0.2000 \text{ L soln} \times \frac{\text{mol NaOH}}{1 \text{ L soln}} \times \frac{\text{mol Cu(OH)}_2}{\text{mol NaOH}} \times \frac{\text{g Cu(OH)}_2}{1 \text{ mol Cu(OH)}_2}$$

$$= \text{ g Cu(OH)}_2$$

$$0.2000 \text{ L soln} \times \frac{0.250 \text{ mol NaOH}}{1 \text{ L soln}} \times \frac{1 \text{ mol Cu(OH)}_2}{2 \text{ mol NaOH}} \times \frac{97.57 \text{ g Cu(OH)}_2}{1 \text{ mol Cu(OH)}_2}$$

$$= 2.44 \text{ g Cu(OH)}_2$$

Where this goes wrong



given: 150.0 mL of 0.300 M KI · correct: 0.1500 L → 0.0450 mol KI → 0.0225 mol PbI₂ → 10.4 g

Feeding milliliters to the molarity. $150.0 \times 0.300 = 45.0$ mol KI, a thousand times too many. Molarity counts liters of solution: 150.0 mL is 0.1500 L, and $0.1500 \times 0.300 = 0.0450$ mol.

Skipping the mole ratio. $0.0450 \times 461.0 = 20.7$ g assumes one PbI₂ per KI. The equation gives 1 PbI₂ : 2 KI, and only the mole ratio switches substances.

Inverting the ratio. (2 mol KI / 1 mol PbI₂) leaves mol KI uncanceled, and $0.0450 \times 2 \times 461.0 = 41.5$ g is wrong. Write the wanted substance on top, so the given unit cancels.

 Practice 1

molar mass Fe(OH)_3 106.87 g/mol

300.0 mL of 0.200 M NaOH reacts completely with excess iron(III) chloride, and rust-brown Fe(OH)_3 precipitates. What mass of Fe(OH)_3 , in grams, forms?

- A. 2.14
- B. 6.41
- C. 19.2
- D. 0.0200

✓ Practice 1: answer A



given: 300.0 mL of 0.200 M NaOH · wanted: g Fe(OH)₃ (106.87 g/mol)

$$0.3000 \text{ L soln} \times \frac{0.200 \text{ mol NaOH}}{1 \text{ L soln}} \times \frac{1 \text{ mol Fe(OH)}_3}{3 \text{ mol NaOH}} \times \frac{106.87 \text{ g Fe(OH)}_3}{1 \text{ mol Fe(OH)}_3}$$

= 2.14 g Fe(OH)₃ (answer A)

B skipped the mole ratio: $0.0600 \times 106.87 = 6.41$. C inverted the ratio: $0.0600 \times 3 \times 106.87 = 19.2$. D stopped at moles: $0.0600 \times (1/3) = 0.0200 \text{ mol Fe(OH)}_3$, one factor short of grams.

 **Sense check:** 0.0600 mol NaOH gives a third as many moles of Fe(OH)₃: 0.0200 mol, at about 107 g per mole: about 2 g. 2.14 g ✓



Worked example 3: volume of a second solution



given: 175.0 mL of 0.240 M NaOH · wanted: mL of 0.150 M H₂SO₄

A base spill of **175.0 mL of 0.240 M NaOH** is neutralized with **0.150 M H₂SO₄** from the shelf. What volume of the acid solution, in milliliters, reacts completely?

Count the factors on the route: mL → L → mol NaOH → mol H₂SO₄ → L → mL.

✓ Worked example 3: solution



Step 1 • Volume → moles

$$0.1750 \text{ L NaOH soln} \times \frac{0.240 \text{ mol NaOH}}{1 \text{ L NaOH soln}} = 0.0420 \text{ mol NaOH}$$

Step 2 • Moles → moles

Step 3 • Moles → the wanted unit

$$0.1750 \text{ L NaOH soln} \times \frac{0.240 \text{ mol NaOH}}{1 \text{ L NaOH soln}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} \times \frac{1 \text{ L acid soln}}{0.150 \text{ mol H}_2\text{SO}_4}$$

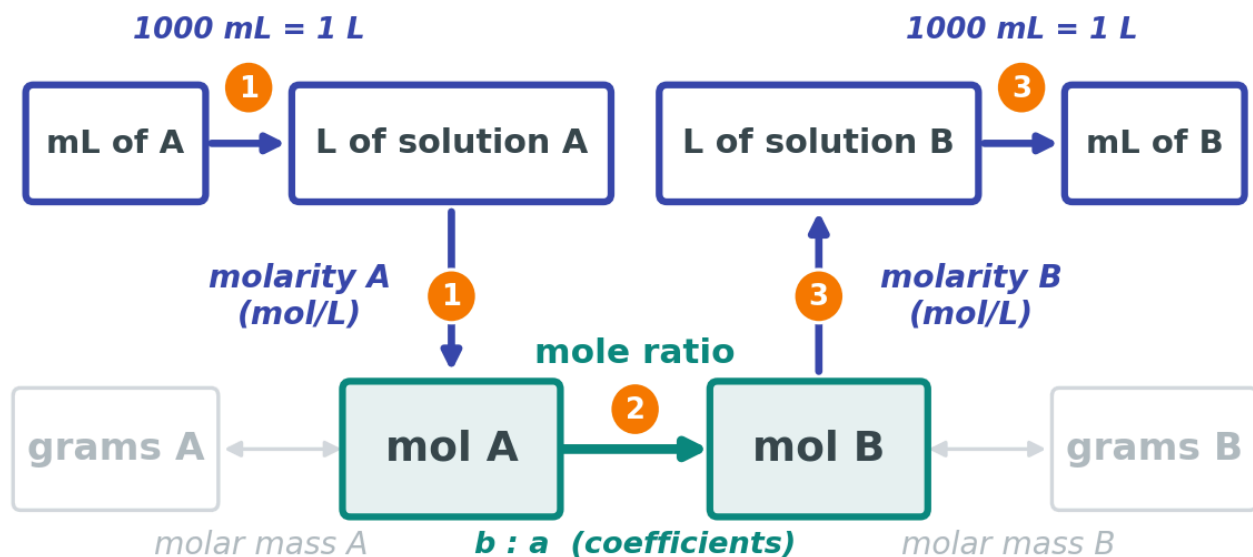
$$= 0.140 \text{ L} = \mathbf{140. \text{ mL acid soln}}$$

🧠 **Sense check:** Half the moles (0.0210 vs 0.0420), but fewer per liter (0.150 vs 0.240): 140. mL is near 175.0 mL ✓

✓ Worked example 3: the route on the map



given: 175.0 mL of 0.240 M NaOH · found: 140. mL of 0.150 M H₂SO₄



🧠 **Sense check:** A is NaOH, B is H₂SO₄. Molarity works at both ends: the base's in, before the mole ratio, and the acid's out, after it. ✓

 Practice 2

two labeled solutions: 0.250 M Ba(OH)_2 · 0.400 M HCl

150.0 mL of 0.250 M Ba(OH)_2 reacts completely with **0.400 M HCl**. What volume of the HCl solution, in milliliters, is required?

- A. 0.0750
- B. 46.9
- C. 93.8
- D. 188

✓ Practice 2: answer D



given: 150.0 mL of 0.250 M Ba(OH)₂ · wanted: mL of 0.400 M HCl

$$0.1500 \text{ L base soln} \times \frac{0.250 \text{ mol Ba(OH)}_2}{1 \text{ L base soln}} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Ba(OH)}_2} \times \frac{1 \text{ L acid soln}}{0.400 \text{ mol HCl}}$$

= 0.1875 L = **188 mL (answer D)**

A stopped at moles: $0.0375 \times (2/1) = 0.0750 \text{ mol HCl}$, one factor short of a volume. B inverted the mole ratio: $0.0375 \times 1/2 \div 0.400 = 0.0469 \text{ L}$, or 46.9 mL. C skipped the mole ratio: $0.0375 \div 0.400 = 0.0938 \text{ L}$, or 93.8 mL.

 **Sense check:** Two HCl are needed per Ba(OH)₂, and the acid is not twice as concentrated (0.400 vs 0.250 M), so the acid volume comes out larger: 188 mL vs 150.0 mL ✓

15 · Quantitative Analysis

Turn one careful measurement of a reaction (a titrant volume, a precipitate mass, trapped CO_2 and H_2O) into the unknown concentration, mass percent, or formula behind it.

The chloride in a water sample



one mass at the end reads the amount at the start

A water sample holds dissolved chloride no balance can weigh. Make it react into a solid, weigh that, and the equation reports what the water held.

A reaction measures what it consumes



weighed: g AgCl on the balance · computed: g Cl⁻ that reacted · the mole ratio links them

A balanced equation ties every amount to every other amount. Measure any one of them, and stoichiometry computes the rest. Quantitative analysis is that strategy: react the unknown, measure something convenient.

The same stoichiometry, started from a measurement



mass-to-mass stoichiometry: molar mass · mole ratio · molar mass

mL of titrant → mol titrant → mol unknown → molarity of the unknown

quantitative analysis: molarity · mole ratio · the sample's own volume

Mass-to-mass problems start from an amount someone chose. Quantitative analysis starts from an amount the lab measured. The mole ratio in the middle does not change.

Three measurements, one skill

titration: volume of a known solution

a buret reads the volume that just completes the reaction · molarity converts it to moles

gravimetric analysis: mass of a precipitate

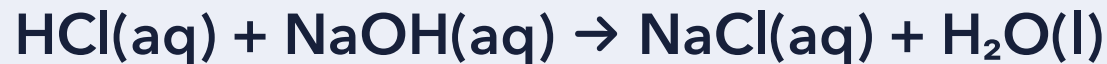
the unknown ion is collected as a weighable solid · molar mass converts it to moles

combustion analysis: masses of CO_2 and H_2O

a burned sample's carbon and hydrogen land in traps · each molar mass converts a gain to moles

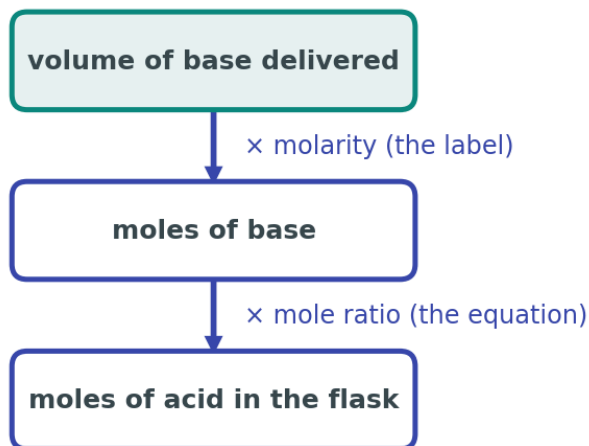
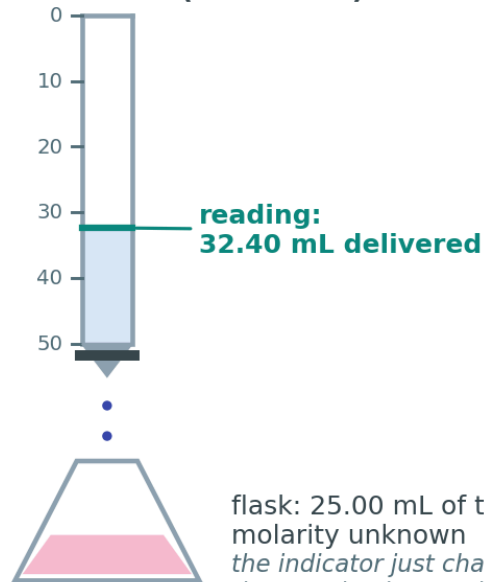
Each technique measures one quantity and converts it to moles. From there every problem is the same stoichiometry.

Titration counts moles with a buret



flask: the acid, amount unknown · buret: NaOH of known molarity (the titrant), volume read to 0.01 mL

buret: 0.125 M NaOH (the titrant)



*one volume reading counts the moles
the flask never told you*

Base flows in until an indicator's color just changes: the reaction is complete. The buret's volume reading becomes moles of acid.

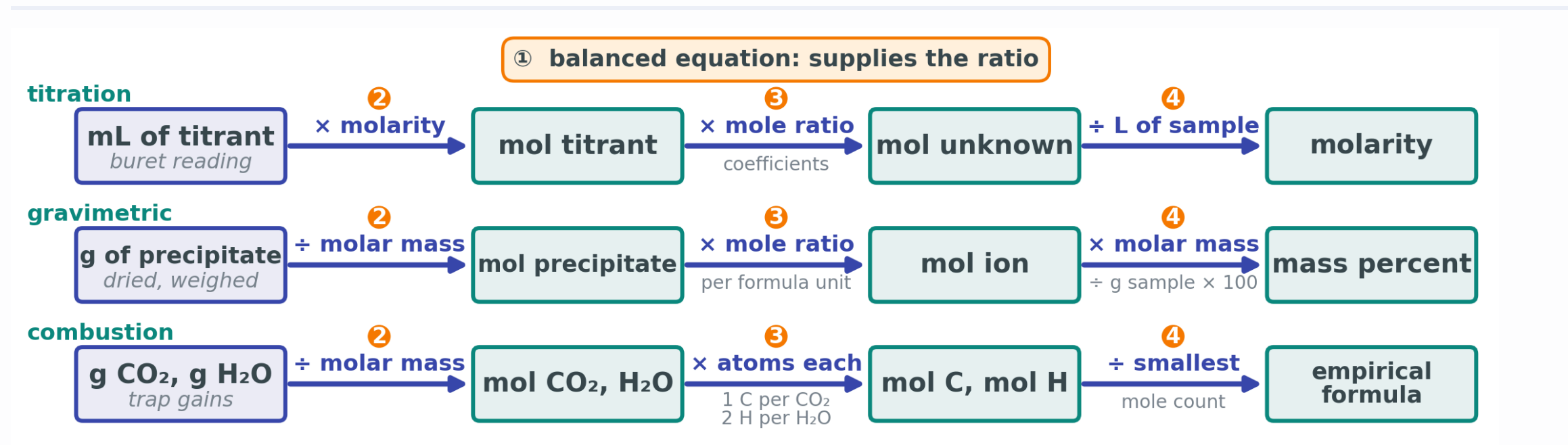


The method

Titration means stoichiometry so always start with a balanced chemical equation.

1. Write the balanced equation.
2. **Convert the measurement to moles:** molar mass or molarity.
3. **Cross to the unknown with the mole ratio.**
4. **Report the asked-for form.**

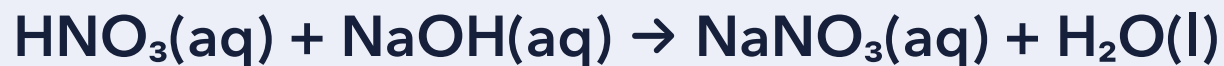
The method on the map



Step 2 depends on what was measured. Step 4 depends on what the question asks. Step 3 is always the mole ratio from the balanced equation.

Guided example: acid in an etching rinse

Step 1 • Write the balanced equation

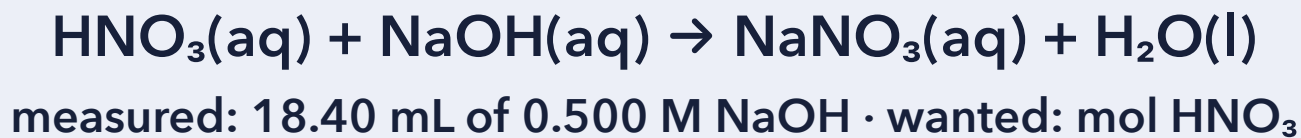


measured: 18.40 mL of 0.500 M NaOH to the color change • wanted: mol HNO_3 in the flask

Rinse water from a metal-etching bath carries nitric acid. A sample is titrated with **0.500 M NaOH**; the indicator changes color at **18.40 mL**. How many moles of HNO_3 were in the flask?

On the map: the titration row, from mL of titrant to mol unknown.

✓ Guided example: solution



Step 2 · Convert the measurement to moles

$$0.01840 \cancel{\text{ L NaOH soln}} \times \frac{0.500 \text{ mol NaOH}}{1 \cancel{\text{ L NaOH soln}}} = 0.00920 \text{ mol NaOH}$$

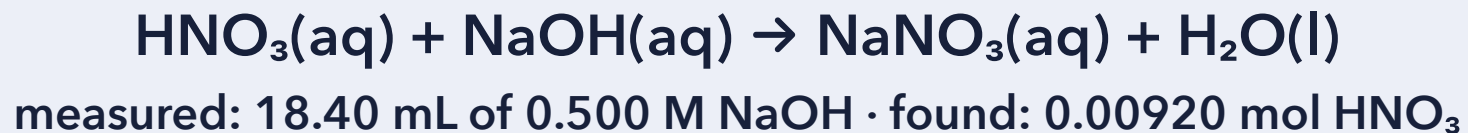
Step 3 · Cross to the unknown with the mole ratio

Step 4 · Report the asked-for form

$$0.00920 \cancel{\text{ mol NaOH}} \times \frac{1 \text{ mol HNO}_3}{1 \cancel{\text{ mol NaOH}}} = 0.00920 \text{ mol HNO}_3$$

 **Sense check:** A fiftieth of a liter (20.00 mL) of 0.500 M base is 0.0100 mol; 18.40 mL is a little less, and one HNO₃ met each NaOH ✓

✓ Guided example: the route on the map



🧠 **Sense check:** The titration row, Steps 2 and 3. The question asked for moles, so the molarity exit stays unused. ✓



Worked example 1: an unknown acid concentration

Step 1 · Write the balanced equation



given: 25.00 mL of HCl solution, molarity unknown · measured: 32.40 mL of 0.125 M NaOH to the color change

A stockroom bottle of hydrochloric acid lost its label. A **25.00 mL** sample is titrated with **0.125 M NaOH**; the indicator changes color at **32.40 mL**. Find the molarity of the acid.

Write the route: mL of base $\rightarrow$ mol NaOH $\rightarrow$ mol HCl $\rightarrow$ M of the acid.

✓ Worked example 1: solution



given: 25.00 mL acid sample · measured: 32.40 mL of 0.125 M NaOH · wanted: M of HCl

Step 2 · Convert the measurement to moles

$$0.03240 \text{ L NaOH soln} \times \frac{0.125 \text{ mol NaOH}}{1 \text{ L NaOH soln}} = 0.00405 \text{ mol NaOH}$$

Step 3 · Cross to the unknown with the mole ratio

Step 4 · Report the asked-for form

$$\frac{0.00405 \text{ mol HCl}}{0.02500 \text{ L acid soln}} = 0.162 \text{ M HCl}$$

 **Sense check:** The acid needed more base volume than its own (32.40 vs 25.00 mL), so it packs more moles per liter than the base: 0.162 M vs 0.125 M ✓

✓ Worked example 1: the route on the map



given: 25.00 mL acid sample · 32.40 mL of 0.125 M NaOH · found: 0.162 M HCl



Sense check: All four steps of the titration row. Step 4 divides by the acid sample's 25.00 mL, never the buret's 32.40 mL. ✓

 Practice 1

given: 20.00 mL of H_2SO_4 solution, molarity unknown · measured: 28.00 mL of 0.150 M NaOH to the color change

A **20.00 mL** sample of diluted battery acid is titrated with **0.150 M NaOH**; the indicator changes color at **28.00 mL**. What is the molarity of the H_2SO_4 ?

- A. 0.420
- B. 0.105
- C. 0.210
- D. 0.00210

✓ Practice 1: answer B



given: 20.00 mL acid sample · measured: 28.00 mL of 0.150 M NaOH · wanted: M of H_2SO_4

$$0.02800 \text{ L NaOH soln} \times \frac{0.150 \text{ mol NaOH}}{1 \text{ L NaOH soln}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} = 0.00210 \text{ mol H}_2\text{SO}_4$$

$$\frac{0.00210 \text{ mol H}_2\text{SO}_4}{0.02000 \text{ L acid soln}} = \mathbf{0.105 \text{ M (answer B)}}$$

A flipped the mole ratio: $0.00840 \div 0.02000 = 0.420 \text{ M}$. C skipped the 2 : 1 ratio: $0.00420 \div 0.02000 = 0.210 \text{ M}$. D stopped at moles: 0.00210 mol H_2SO_4 , never divided by the 0.02000 L sample.

🧠 **Sense check:** Two NaOH per H_2SO_4 : the acid holds half the base's moles, so 0.105 M sits below 0.150 M ✓

✓ Practice 1: the route on the map



given: 20.00 mL acid sample · 28.00 mL of 0.150 M NaOH · found: 0.105 M H₂SO₄



Sense check: The titration row, all four steps. Step 3 carries the equation's 1 H₂SO₄ : 2 NaOH, so the acid holds half the base's moles. ✓

Gravimetric analysis: collect the ion as a solid



excess Ag^+ leaves essentially no Cl^- behind in solution · the solid is filtered, dried, weighed

Excess reagent converts the ion completely into one insoluble solid. The dried precipitate's mass is the measurement; molar mass and mole ratio read the ion back out of it.

Worked example 2: chloride in a de-icing blend

Step 1 • Write the balanced equation



given: 0.6250 g sample, dissolved • measured: dried AgCl 1.0260 g • wanted: mass % $\text{Cl}^- = \text{g Cl}^- \div \text{g sample} \times 100$

A **0.6250 g** sample of a road de-icing blend is dissolved in water; excess AgNO_3 precipitates the chloride. The dried AgCl weighs **1.0260 g**. Find the mass percent of chloride in the blend. (*AgCl 143.32 g/mol • Cl 35.45 g/mol*)

The solid outweighs the sample it came from. The chloride percent must still land below 100.

✓ Worked example 2: grams of chloride



given: 0.6250 g sample · measured: 1.0260 g AgCl · wanted: mass % Cl⁻

Step 2 · Convert the measurement to moles

$$1.0260 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.32 \text{ g AgCl}} = 0.00716 \text{ mol AgCl}$$

Step 3 · Cross to the unknown with the mole ratio

$$1.0260 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.32 \text{ g AgCl}} \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol AgCl}} \times \frac{35.45 \text{ g Cl}^-}{1 \text{ mol Cl}^-} = 0.2538 \text{ g Cl}^-$$

 **Sense check:** 0.2538 g of chloride from 1.0260 g of solid: most of the precipitate's mass is the silver ✓

✓ Worked example 2: the percent

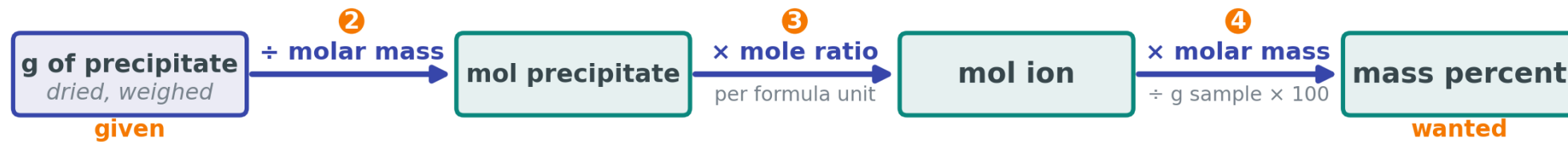


given: 0.6250 g sample · found: 0.2538 g Cl^- in the 1.0260 g of AgCl

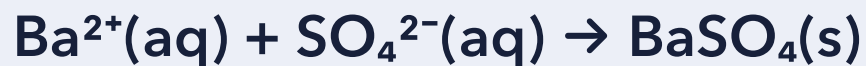
Step 4 · Report the asked-for form

$$\frac{0.2538 \text{ g Cl}^-}{0.6250 \text{ g sample}} \times 100 = 40.6\% \text{ Cl}^-$$

 **Sense check:** Below 100 ✓. Only 35.45 of every 143.32 g of AgCl is chloride, so 1.0260 g of solid holds just 0.2538 g of Cl^- . Pure NaCl would be 60.7% chloride; this blend, with non-chloride components, lands lower.



Your turn: sulfate in a fertilizer



given: 0.8000 g fertilizer, dissolved · measured: dried BaSO_4 0.5580 g · wanted: mass % SO_4^{2-}

$$0.5580 \text{ g BaSO}_4 \times \frac{1 \text{ mol BaSO}_4}{\text{g BaSO}_4} \times \frac{1 \text{ mol SO}_4^{2-}}{1 \text{ mol BaSO}_4} \times \frac{\text{g SO}_4^{2-}}{1 \text{ mol SO}_4^{2-}} = \text{g SO}_4^{2-}$$

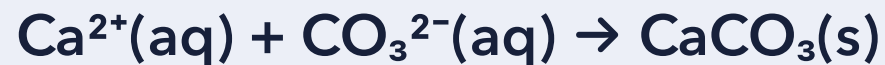
$$\frac{\text{g SO}_4^{2-}}{0.8000 \text{ g sample}} \times 100 = \%$$

$$0.5580 \text{ g BaSO}_4 \times \frac{1 \text{ mol BaSO}_4}{233.39 \text{ g BaSO}_4} \times \frac{1 \text{ mol SO}_4^{2-}}{1 \text{ mol BaSO}_4} \times \frac{96.06 \text{ g SO}_4^{2-}}{1 \text{ mol SO}_4^{2-}}$$

$$= 0.2297 \text{ g SO}_4^{2-} \cdot \frac{0.2297 \text{ g SO}_4^{2-}}{0.8000 \text{ g sample}} \times 100 = 28.7\%$$



Practice 2

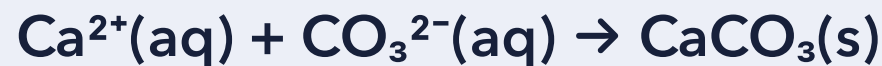


given: 1.2500 g sample, dissolved · measured: dried CaCO_3 0.6900 g · CaCO_3 100.09 g/mol · Ca 40.08 g/mol

A **1.2500 g** sample of a mineral supplement is dissolved, and excess sodium carbonate precipitates the calcium as CaCO_3 . The dried solid weighs **0.6900 g**. What is the mass percent of calcium in the sample?

- A. 55.20
- B. 22.10
- C. 137.8
- D. 0.2763
- E. 40.04

✓ Practice 2: answer B

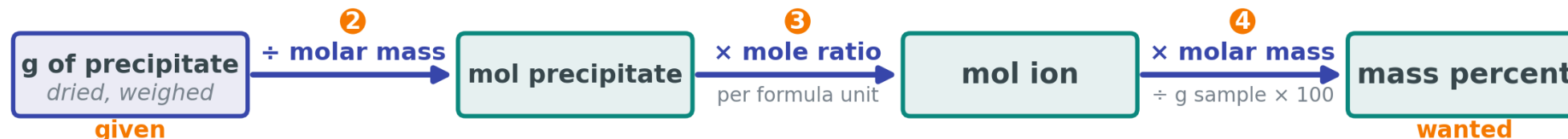


given: 1.2500 g sample · measured: 0.6900 g CaCO_3 · wanted: mass % Ca^{2+}

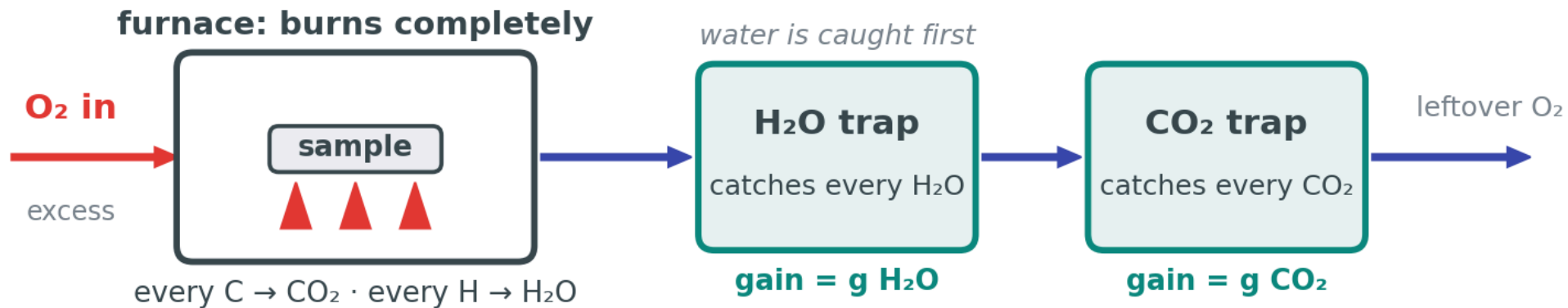
$$0.6900 \text{ g } \text{CaCO}_3 \times \frac{1 \text{ mol } \text{CaCO}_3}{100.09 \text{ g } \text{CaCO}_3} \times \frac{1 \text{ mol } \text{Ca}^{2+}}{1 \text{ mol } \text{CaCO}_3} \times \frac{40.08 \text{ g } \text{Ca}^{2+}}{1 \text{ mol } \text{Ca}^{2+}} = 0.2763 \text{ g } \text{Ca}^{2+}$$

$$\cdot \frac{0.2763 \text{ g } \text{Ca}^{2+}}{1.2500 \text{ g sample}} \times 100 = \mathbf{22.10\% \text{ (answer B)}}$$

🧠 **Sense check:** CaCO_3 is 40.04% calcium, so the 0.6900 g of solid holds 0.2763 g of it, about a fifth of the 1.2500 g sample ✓



Combustion analysis: traps sort the atoms



each trap is weighed before and after: its gain is one element's measurement

Burned in excess oxygen, a compound of carbon and hydrogen releases every C atom as CO₂ and every H atom in H₂O. Each trap's gain in mass measures one element.

Worked example 3: a hydrocarbon's empirical formula

Step 1 • Write the balanced equation

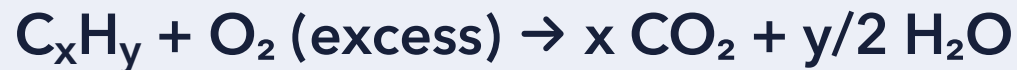


measured: CO_2 trap +0.8800 g · H_2O trap +0.3604 g · sample burned: 0.2805 g · wanted: empirical formula

A **0.2805 g** sample of a fuel gas containing only carbon and hydrogen burns completely. The CO_2 trap gains **0.8800 g**; the H_2O trap gains **0.3604 g**. Find the empirical formula.
(CO_2 44.01 · H_2O 18.02 g/mol)

Write the route: g $\text{CO}_2 \rightarrow$ mol C, g $\text{H}_2\text{O} \rightarrow$ mol H, then the smallest whole-number ratio.

✓ Worked example 3: moles of C and H



measured: +0.8800 g CO₂ · +0.3604 g H₂O · sample: 0.2805 g · wanted: empirical formula

Step 2 · Convert the measurement to moles

Step 3 · Cross to the unknown with the mole ratio

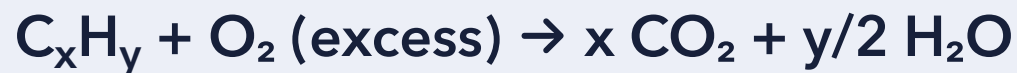
$$0.8800 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.02000 \text{ mol C}$$

The hydrogen chain

$$0.3604 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.04000 \text{ mol H}$$

 **Sense check:** 0.02000 mol C and 0.04000 mol H: twice as many hydrogen atoms as carbon ✓

✓ Worked example 3: the formula finish

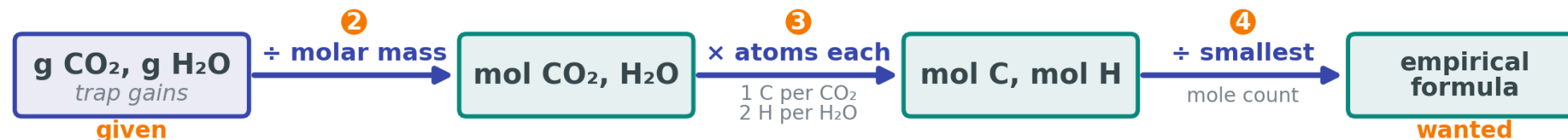


found: 0.02000 mol C · 0.04000 mol H · sample: 0.2805 g · wanted: empirical formula

Step 4 · Report the asked-for form

$$\text{C: } \frac{0.02000 \text{ mol}}{0.02000} = 1.00 \cdot \text{H: } \frac{0.04000 \text{ mol}}{0.02000} = 2.00 = \text{CH}_2$$

 **Sense check:** Mass check: 0.02000 mol C weighs 0.2402 g, 0.04000 mol H weighs 0.0403 g; together 0.2805 g, the whole sample, so no other element hides in it ✓. Here the division came out 2.00; a ratio that lands on 1.33 is the fraction 4/3, and multiplying every count by 3 clears it, the same multiply-to-whole finish that turns percent-composition data into formulas. A measured molar mass of 28.05 g/mol would scale CH₂ to C₂H₄.



⚠ Where this goes wrong

Reporting the precipitate's percent. 1.0260 g of AgCl from a 0.6250 g sample reads $1.0260 \div 0.6250 \times 100 = 164.2\%$, more chloride than sample. The silver's mass came from the added reagent. Only 35.45 of every 143.32 g of AgCl is chloride: 0.2538 g, and 40.6%.

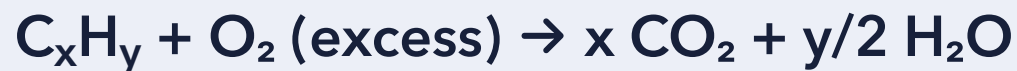
Taking a trap's gain as the element's mass. The CO₂ trap gained 0.8800 g, but that mass is carbon plus oxygen. Carbon is 12.01 of every 44.01 g: $0.8800 \times 12.01 \div 44.01 = 0.2402$ g C.

Forgetting each H₂O carries two H. 0.3604 g of H₂O is 0.02000 mol H₂O but 0.04000 mol H. Counting 0.02000 mol H turns CH₂ into CH, a different compound.

Dividing by the titrant's volume. 0.00405 mol HCl $\div$ 0.03240 L gives 0.125 M, the base's own molarity back again. The unknown's molarity divides by the unknown's volume: $0.00405 \div 0.02500 = 0.162$ M.



Practice 3

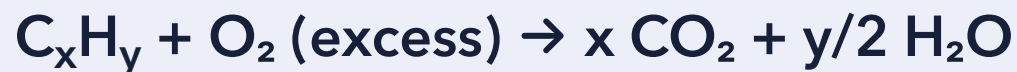


measured: CO_2 trap +1.0562 g · H_2O trap +0.5766 g · CO_2 44.01 · H_2O 18.02 g/mol

A camp-stove fuel contains only carbon and hydrogen. Complete combustion of a sample sends **1.0562 g** into the CO_2 trap and **0.5766 g** into the H_2O trap. What is the empirical formula?

- A. C_3H_8
- B. CH_3
- C. C_3H_4
- D. C_8H_3

✓ Practice 3: answer A



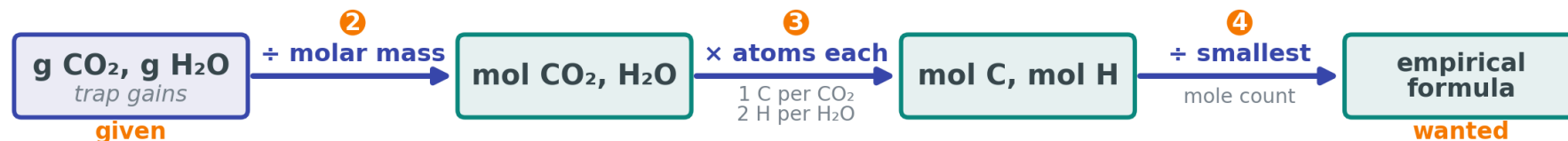
measured: +1.0562 g CO₂ · +0.5766 g H₂O · wanted: empirical formula

$$1.0562 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.02400 \text{ mol C} \cdot 0.5766 \text{ g H}_2\text{O} \times$$

$$\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.06400 \text{ mol H} \cdot \text{C: } 0.02400 \div 0.02400 = 1.00 \cdot$$

$$\text{H: } 0.06400 \div 0.02400 = 2.67 \cdot \times 3 \rightarrow \text{C } 3.00, \text{H } 8.00 = \text{C}_3\text{H}_8 (\text{answer A})$$

🧠 **Sense check:** 2.67 is the fraction $\frac{8}{3}$, and $\times 3$ clears it. C₃H₈ is propane, a real camp-stove fuel ✓





Check yourself

1. A dissolved sample gives 0.4120 g of dried AgCl. State the steps that turn that one mass into the mass percent of chloride in the sample. Which factor comes from the balanced equation?
2. A burned hydrocarbon gives 0.030 mol of CO₂ and 0.020 mol of H₂O. Which element has more moles of atoms in the compound? Compute both counts.

Titration returns with the study of solutions: the exact moment the indicator flips is the equivalence point, and acids that release two or three H⁺ per formula unit change the mole ratio. The divide-by-smallest, multiply-to-whole finish is the same one that turns percent-composition data into formulas.

Can you...?

- ☐ balance any chemical equation with the smallest whole-number coefficients?
- ☐ read a balanced equation as a recipe: coefficients are mole relationships?
- ☐ convert between amounts of any two species using mole ratios?
- ☐ chain molar masses with mole ratios to solve gram-to-gram problems?
- ☐ identify which reactant limits a reaction and how much product it allows?
- ☐ classify a reaction by type and cite the evidence that a reaction occurred?
- ☐ apply the solubility rules to predict precipitates and assign physical states?
- ☐ write molecular, complete ionic, and net ionic equations and classify electrolytes?
- ☐ assign oxidation numbers and identify oxidizing and reducing agents?
- ☐ write half-reactions and use the activity series to predict single displacement?
- ☐ carry molarity through stoichiometric calculations and compute percent yield?
- ☐ find an unknown concentration, mass percent, or formula from one measurement?

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