

Stoichiometry

Preparation for General Chemistry · Dr. Karmach

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

- Read the coefficients of a balanced equation as mole ratios and convert moles of one substance to moles of another
- Convert a mass of one substance to the mass, moles, or number of particles of another through the mole ratio
- Identify the limiting reactant, the maximum product it allows, and the excess reactant left over
- Calculate percent yield from the theoretical and actual yields

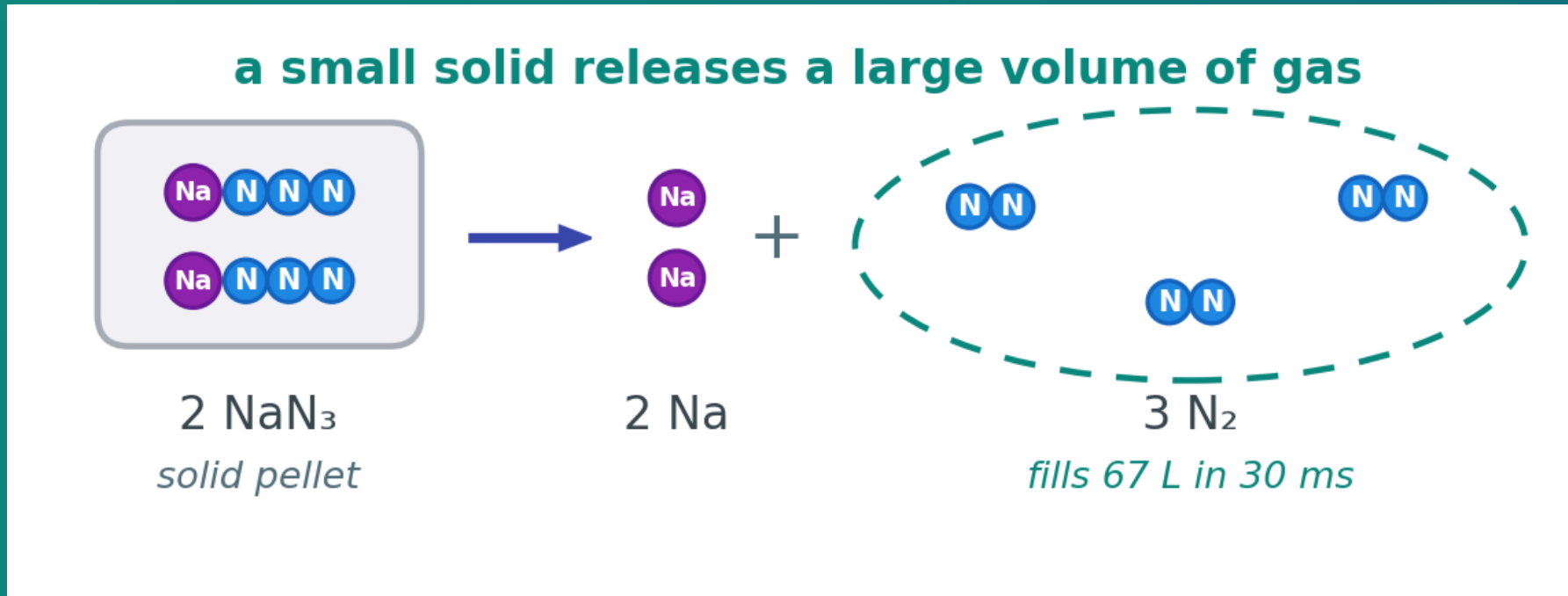
Today's route

1. Mole Ratios
2. Mass-to-Mass Stoichiometry
3. Limiting Reactant
4. Percent Yield

1 • 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.

A balanced equation is a recipe

2 graham crackers + 1 marshmallow + 3 chocolate pieces → 1 s'more

2 : 1 : 3 : 1, fixed by the recipe

A recipe fixes the ratio. Each s'more takes 3 chocolate pieces:

$$12 \text{ ~~pieces~~} \times \frac{1 \text{ s'more}}{3 \text{ ~~pieces~~}} = 4 \text{ s'mores}$$

How many pieces do 10 graham crackers need?

$$10 \text{ ~~crackers~~} \times \frac{3 \text{ pieces}}{2 \text{ ~~crackers~~}} = 15 \text{ pieces}$$

An equation's coefficients work the same way, in moles.

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

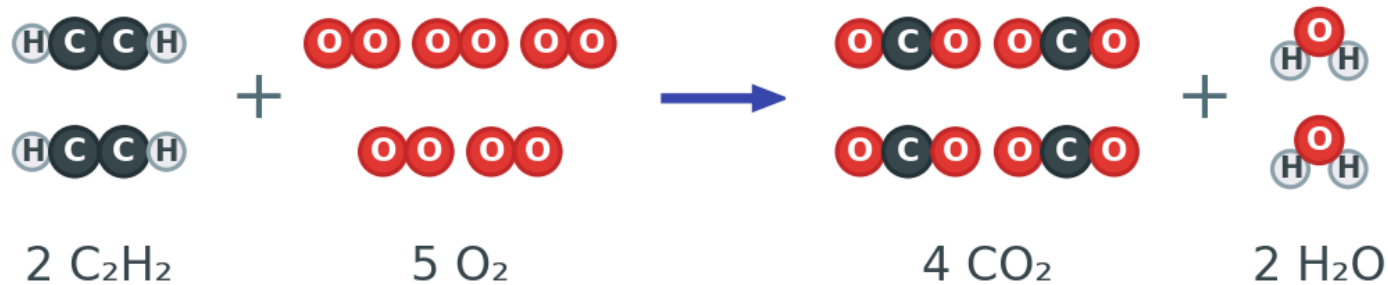


the coefficient 2 may change during balancing ✓ · changing the subscript ₂ would name a different substance ✗

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.

Reading a mole ratio



2 mol H₂ : 1 mol O₂ : 2 mol H₂O · no written coefficient means 1

Pick two substances. Their coefficients, in moles, form a mole ratio, written either way up.

$$\frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} \text{ or } \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2\text{O}} \cdot \frac{2 \text{ mol H}_2}{1 \text{ mol O}_2} \text{ or } \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2}$$

Subscripts never enter a mole ratio. The 2 in H₂O counts hydrogen atoms inside one molecule.

Coefficients become conversion factors



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

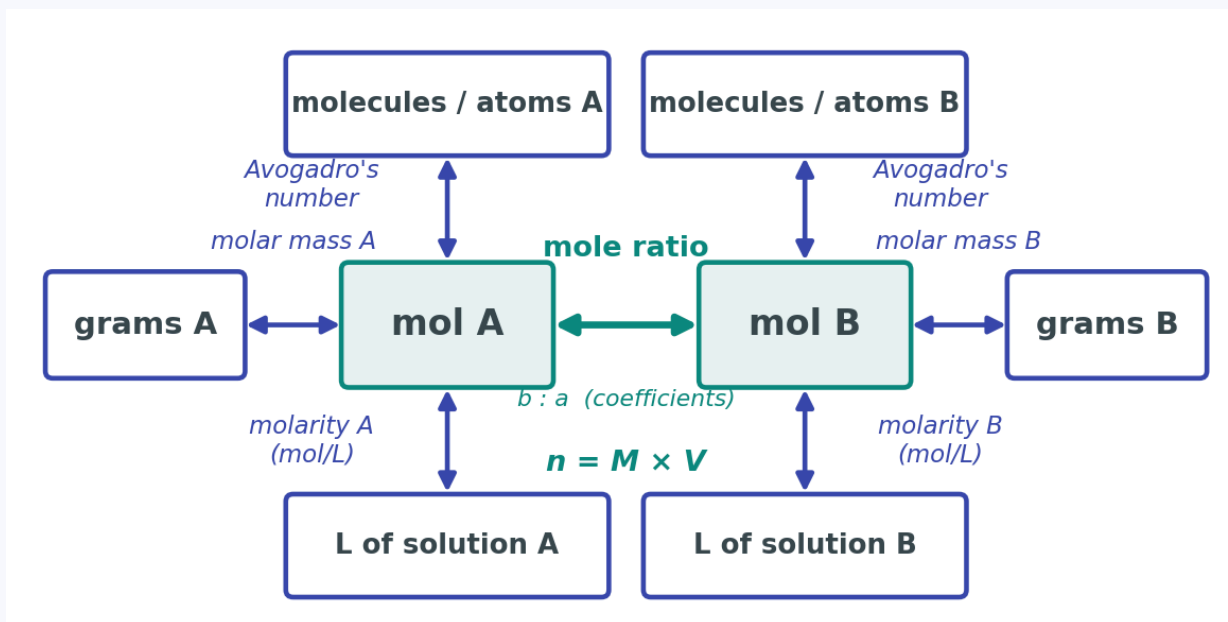
$$\frac{4 \text{ mol CO}_2}{5 \text{ mol O}_2} \text{ or } \frac{5 \text{ mol O}_2}{2 \text{ mol C}_2\text{H}_2} \text{ or any pair}$$

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.





Guided example: a hand warmer

Step 1 · Start from a balanced equation



given: 8.0 mol Fe · wanted: mol O₂

An air-activated hand warmer heats up as iron powder reacts with oxygen. How many moles of O₂ react with **8.0 mol of Fe**?

Read the recipe first: 4 mol Fe take 3 mol O₂.

✓ Guided example: solution



given: 8.0 mol Fe · wanted: mol O₂

Step 2 · Build the ratio, wanted over given

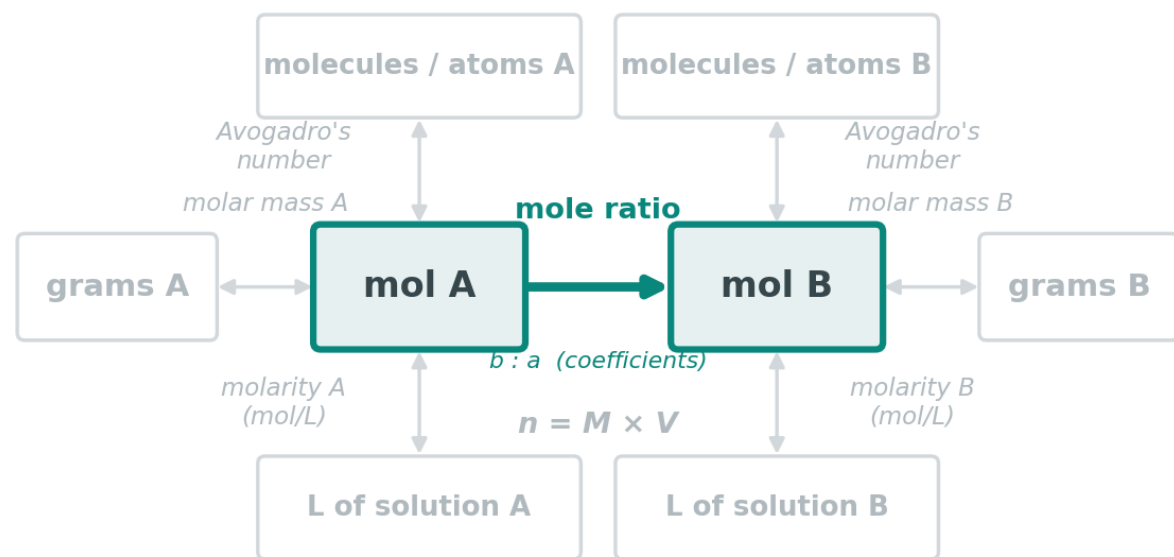
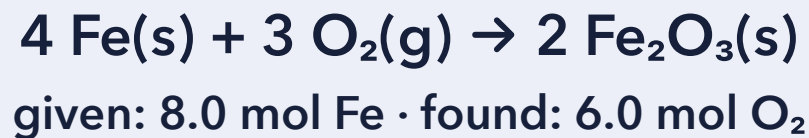
$$\frac{3 \text{ mol O}_2}{4 \text{ mol Fe}} \text{ wanted on top, given on the bottom}$$

Step 3 · Multiply and cancel

$$8.0 \text{ mol Fe} \times \frac{3 \text{ mol O}_2}{4 \text{ mol Fe}} = 6.0 \text{ mol O}_2$$

 **Sense check:** Recipe: 8.0 mol Fe is twice 4, so twice 3 = 6.0 mol O₂. The conversion factor agrees ✓

✓ Guided example: the route on the map



🧠 **Sense check:** A is Fe, B is O₂. A moles-to-moles problem uses only the bridge: one arrow, one mole ratio. ✓

Practice 1: reading the ratio



How many moles of O_2 react with **1 mol of C_2H_6** ?

- A. 3.5
- B. 7
- C. 0.286
- D. 2

✓ Practice 1: answer A



given: 1 mol C₂H₆ · wanted: mol O₂

$$1 \text{ mol C}_2\text{H}_6 \times \frac{7 \text{ mol O}_2}{2 \text{ mol C}_2\text{H}_6} = 3.5 \text{ mol O}_2 \text{ (answer A)}$$

B read O₂'s coefficient alone: 7 mol O₂ pairs with 2 mol C₂H₆, not 1. C inverted the ratio: 1 × (2/7) = 0.286. D answered with C₂H₆'s coefficient, 2, without applying the ratio.

 **Sense check:** The equation pairs 7 O₂ with 2 C₂H₆, so one C₂H₆ takes half of 7. ✓



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

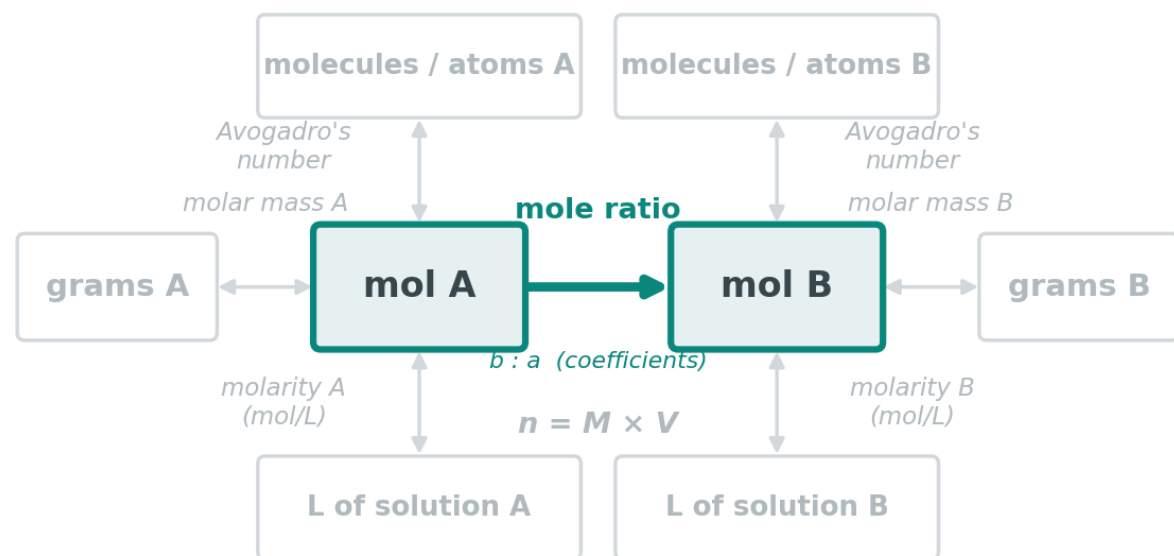
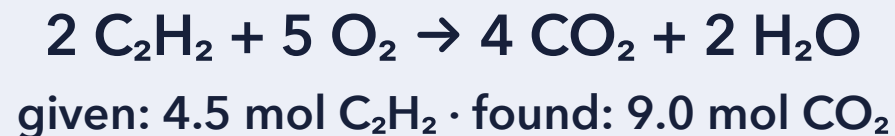
$$\frac{4 \text{ mol CO}_2}{2 \text{ mol C}_2\text{H}_2} \text{ cancels mol C}_2\text{H}_2 \checkmark \quad \frac{2 \text{ mol C}_2\text{H}_2}{4 \text{ mol CO}_2} \text{ cancels nothing } \times$$

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** :

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

Complete the ratio so mol C_2H_2 cancels, then compute.

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

 **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 2

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 2: 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. ✓

Practice 3: oxygen for benzene



A lab burner consumes **0.52 mol of benzene**, C_6H_6 . How many moles of O_2 does it use up?

- A. 0.52
- B. 0.069
- C. 3.1
- D. 3.9

✓ Practice 3: answer D



given: 0.52 mol C₆H₆ · wanted: mol O₂

$$0.52 \cancel{\text{ mol C}_6\text{H}_6} \times \frac{15 \text{ mol O}_2}{2 \cancel{\text{ mol C}_6\text{H}_6}} = 3.9 \text{ mol O}_2 \text{ (answer D)}$$

A assumed 1:1: 0.52. B inverted the ratio: $0.52 \times (2/15) = 0.069$. C read CO₂'s coefficient instead of O₂'s: $0.52 \times (12/2) = 3.1$.

🧠 **Sense check:** Each C₆H₆ takes $15/2 = 7.5$ O₂, so far more O₂ reacts than benzene: $3.9 > 0.52$ ✓

Worked example 2: moles of reactant needed

Step 1 • Start from a balanced equation

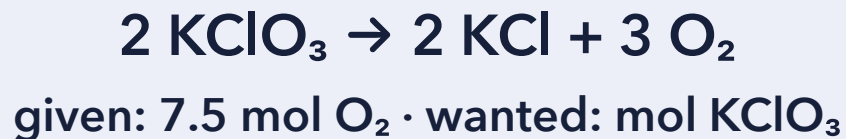


K: $2 = 2 \checkmark$ • Cl: $2 = 2 \checkmark$ • O: $6 = 6 \checkmark$ • given: 7.5 mol O_2 • wanted: mol KClO_3

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



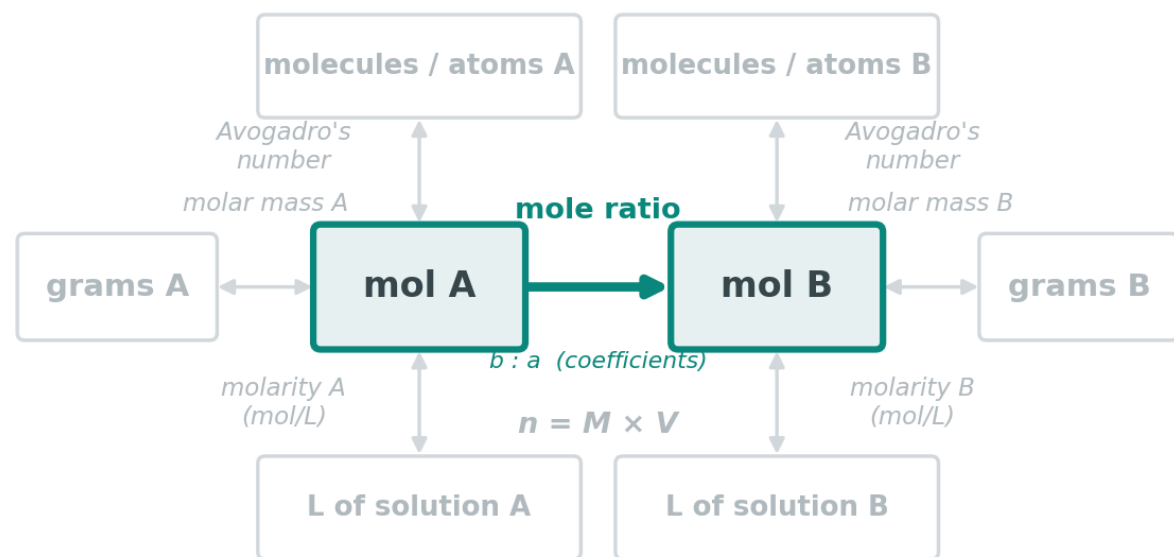
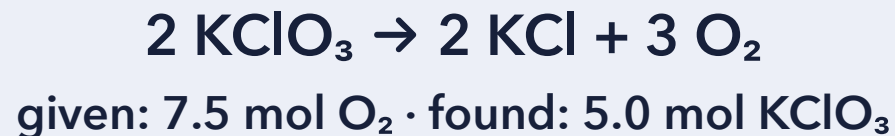
Step 2 · Build the ratio, wanted over given

Step 3 · Multiply and cancel

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

 **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 4: product to product



An ethanol burner releases **0.250 mol of CO₂**. How many moles of H₂O form?

- A. 0.375
- B. 0.250
- C. 0.125
- D. 0.167

✓ Practice 4: answer A



given: 0.250 mol CO₂ · wanted: mol H₂O

$$0.250 \text{ mol CO}_2 \times \frac{3 \text{ mol H}_2\text{O}}{2 \text{ mol CO}_2} = 0.375 \text{ mol H}_2\text{O (answer A)}$$

B assumed 1:1: 0.250. C divided by CO₂'s 2 but left out H₂O's 3: $0.250 \times (1/2) = 0.125$. D inverted the ratio: $0.250 \times (2/3) = 0.167$.

 **Sense check:** The mole ratio links any two substances in the equation, products included. 3 H₂O form for every 2 CO₂: $0.375 > 0.250$ ✓

 Practice 5

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 5: answer C



given: 26.0 mol CO₂ · wanted: mol C₄H₁₀

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

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 CO₂ per butane, so the fuel needed is $26.0 \div 4 = 6.50 \text{ mol}$. ✓

 Practice 6

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 6: answer D



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

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

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

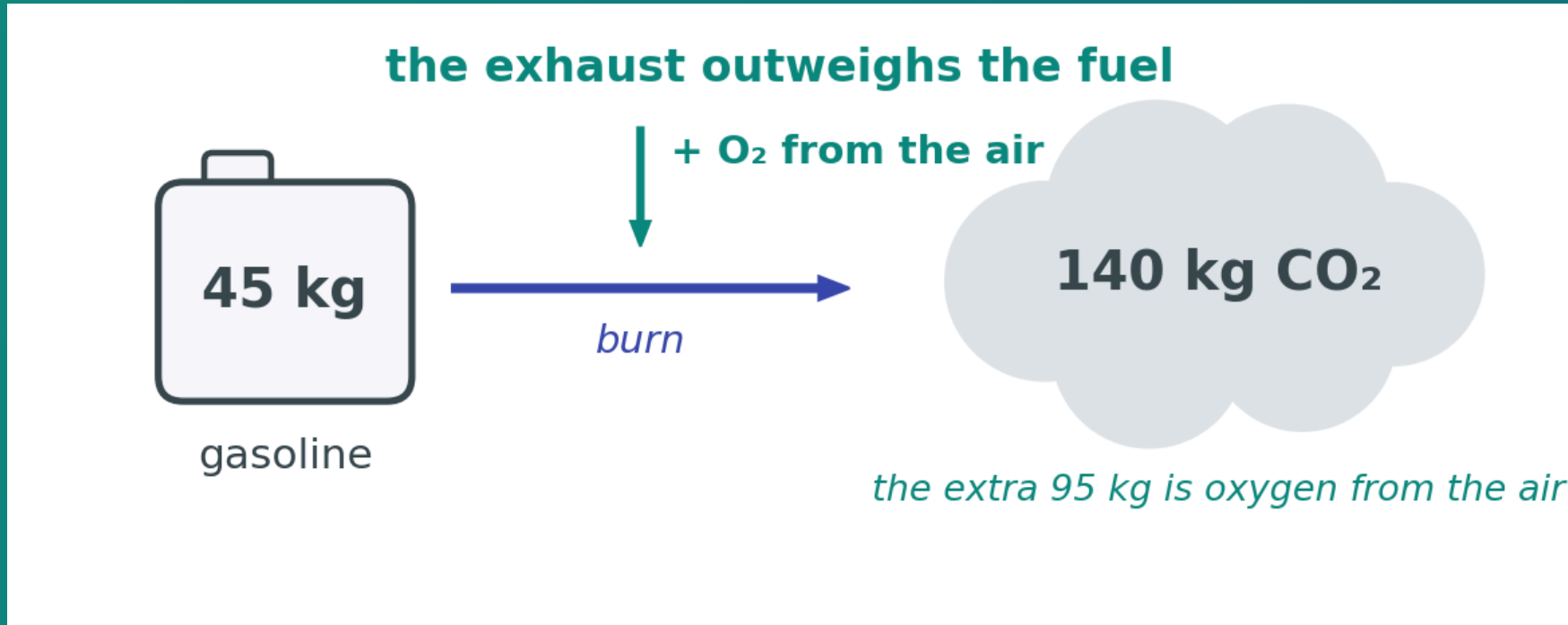
A stopped at the CO₂ alone: 1.50. B assumed 1:1 for each product: 2.40 + 2.40 = 4.80. C inverted both ratios: 2.40 × (8/5) + 2.40 × (8/6) = 3.84 + 3.20 = 7.04.

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

2 · 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

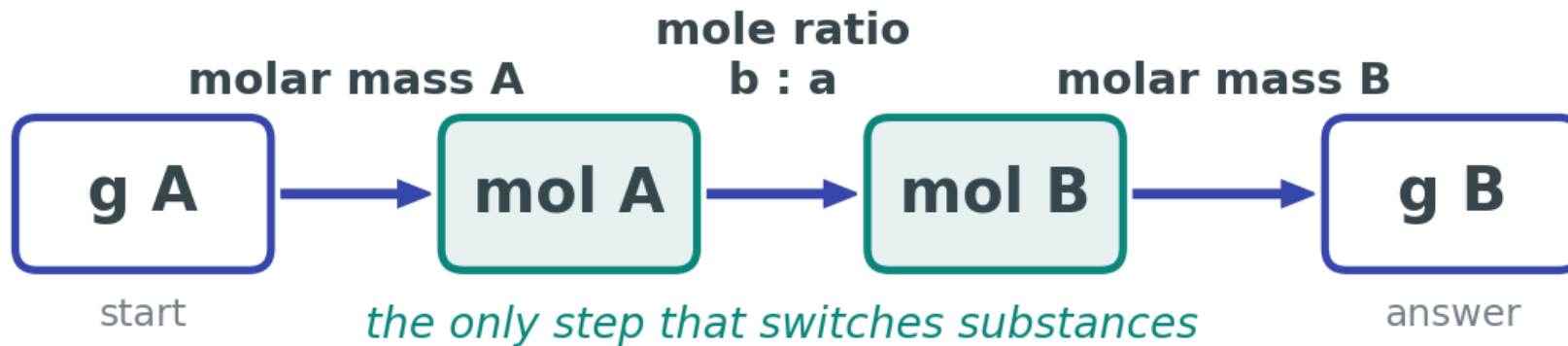


moles = 1 glucose : 6 CO₂ · grams = 180.16 : 264.06, not 1 : 6

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

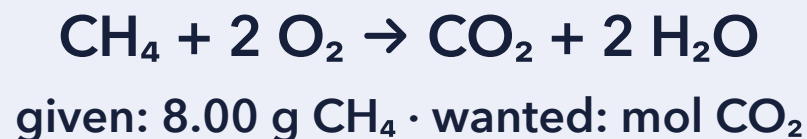


given: 8.00 g CH₄ · wanted: mol CO₂

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

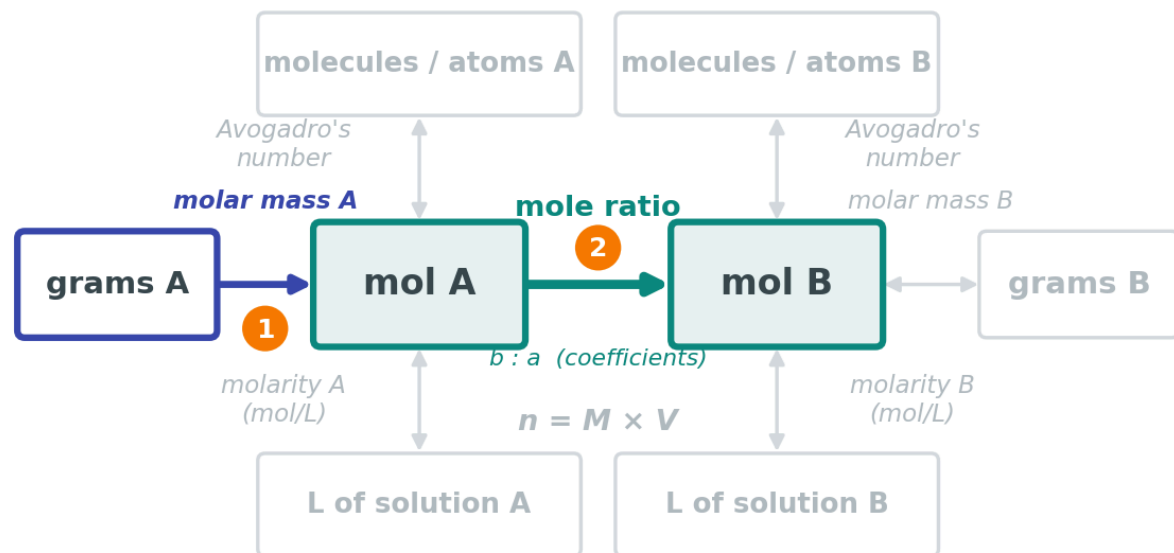
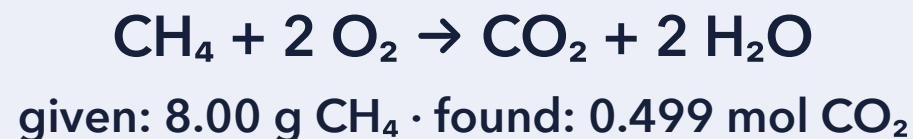
$$8.00 \text{ g } \cancel{\text{CH}_4} \times \frac{1 \text{ mol } \cancel{\text{CH}_4}}{16.04 \text{ g } \cancel{\text{CH}_4}} = 0.499 \text{ mol CH}_4$$

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)

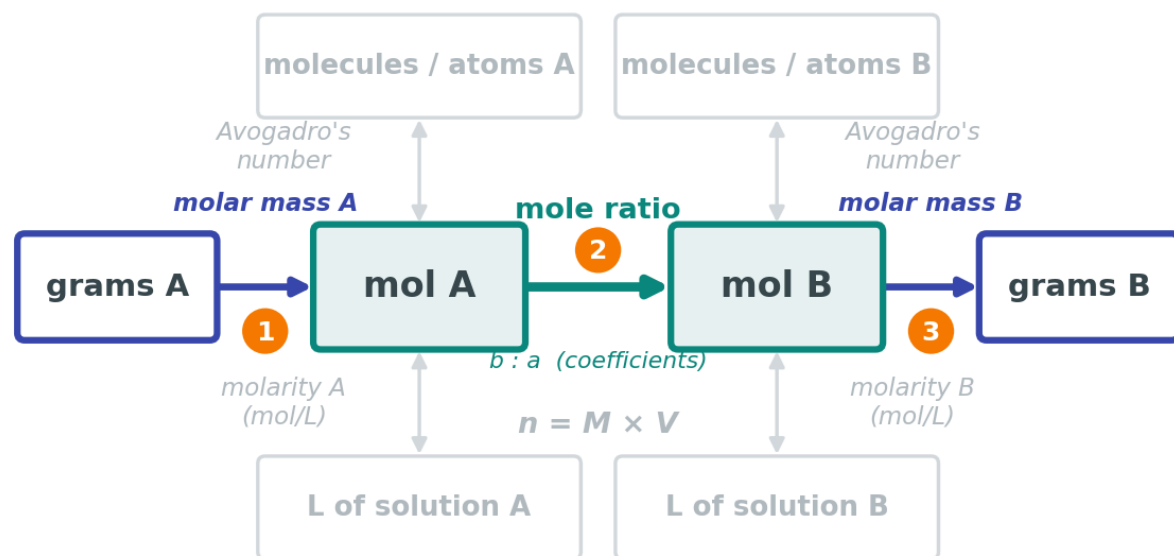
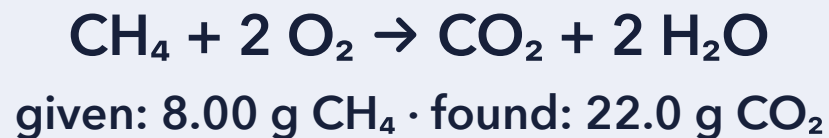
Step 3 · Moles → grams

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

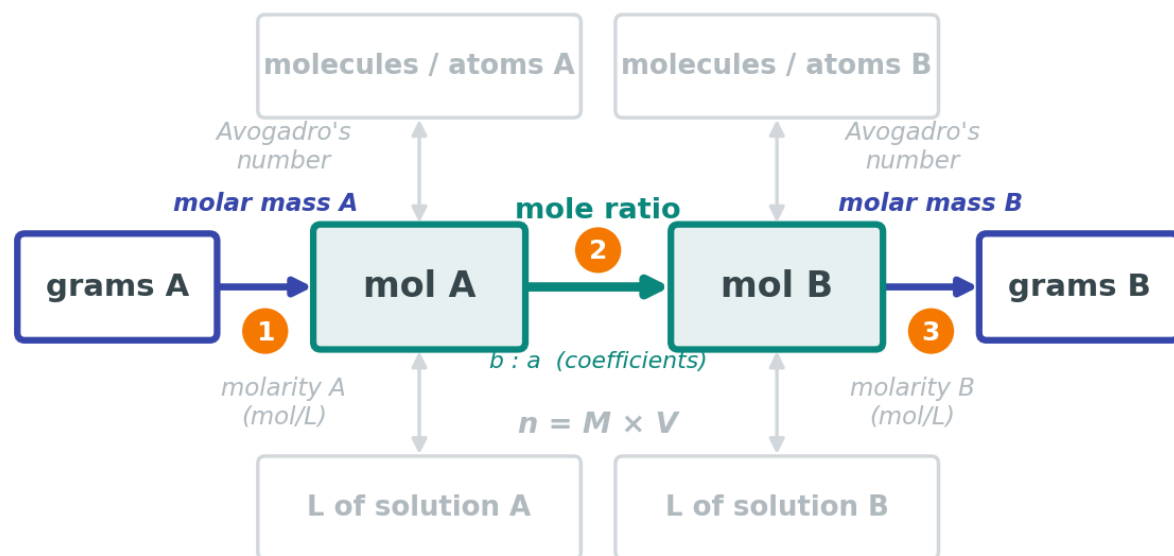
$$90.0 \text{ g glucose} \times \frac{1 \text{ mol glucose}}{180.16 \text{ g glucose}} \times \frac{6 \text{ mol CO}_2}{1 \text{ mol glucose}} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 132 \text{ g CO}_2$$

 **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) reacts completely with oxygen. 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 g NaHCO₃ (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, $2.20 \rightarrow 8.40$ ✓

Worked example 3: grams to molecules



given: 50.0 g NaN_3 · wanted: molecules of N_2

An airbag fills in about 30 milliseconds with N_2 from a pellet of sodium azide. How many **molecules** of N_2 form when **50.0 g of NaN_3** decomposes? (NaN_3 65.02 g/mol)

Write the route first: g NaN_3 $\rightarrow$ mol NaN_3 $\rightarrow$ mol N_2 $\rightarrow$ molecules N_2 . The first two arrows are the grams-to-grams route; only the last factor changes.

✓ Worked example 3: solution



given: 50.0 g NaN₃ · wanted: molecules of N₂

Step 1 · Grams → moles

Step 2 · Moles → moles

Step 3 · Moles → molecules

$$50.0 \text{ g NaN}_3 \times \frac{1 \text{ mol NaN}_3}{65.02 \text{ g NaN}_3} \times \frac{3 \text{ mol N}_2}{2 \text{ mol NaN}_3} \times \frac{6.022 \times 10^{23} \text{ molecules N}_2}{1 \text{ mol N}_2}$$

$$= 6.95 \times 10^{23} \text{ molecules N}_2$$

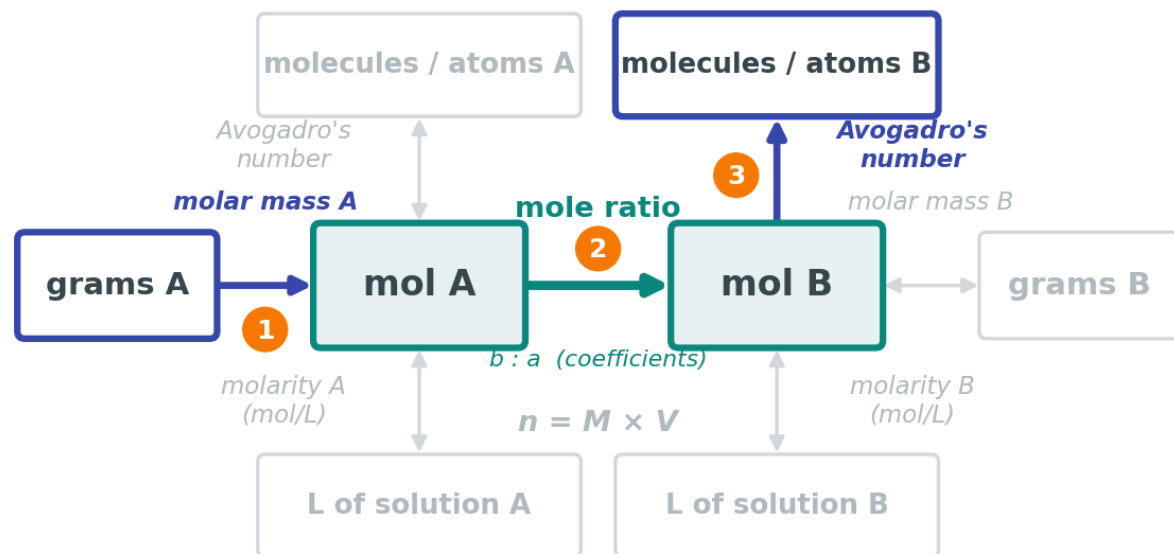
 **Sense check:** Coefficients count particles, so the chain can end on a count as easily as on grams. 1.15 mol is a little over one mole, and the count lands a little over 6.022×10^{23} ✓

Run backward, the chain starts at a count: molecules B → mol B → mol A → g A, with Avogadro's number flipped so molecules cancel.

✓ Worked example 3: the route on the map



given: 50.0 g NaN_3 · found: 6.95×10^{23} molecules N_2



🧠 **Sense check:** Three arrows, three conversion factors: NaN_3 's molar mass, the mole ratio, then Avogadro's number.



 Practice 4

A car engine burns **5.00 g of octane** (C_8H_{18} , 114.22 g/mol). How many molecules of CO_2 form?

- A. 0.350
- B. 2.11×10^{23}
- C. 5.82×10^{-25}
- D. 2.64×10^{22}

✓ Practice 4: answer B



given: 5.00 g C_8H_{18} · wanted: molecules CO_2

$$5.00 \text{ g } \text{C}_8\text{H}_{18} \times \frac{1 \text{ mol } \text{C}_8\text{H}_{18}}{114.22 \text{ g } \text{C}_8\text{H}_{18}} \times \frac{16 \text{ mol } \text{CO}_2}{2 \text{ mol } \text{C}_8\text{H}_{18}} \times \frac{6.022 \times 10^{23} \text{ molecules } \text{CO}_2}{1 \text{ mol } \text{CO}_2}$$

= 2.11×10^{23} molecules CO_2 (answer B)

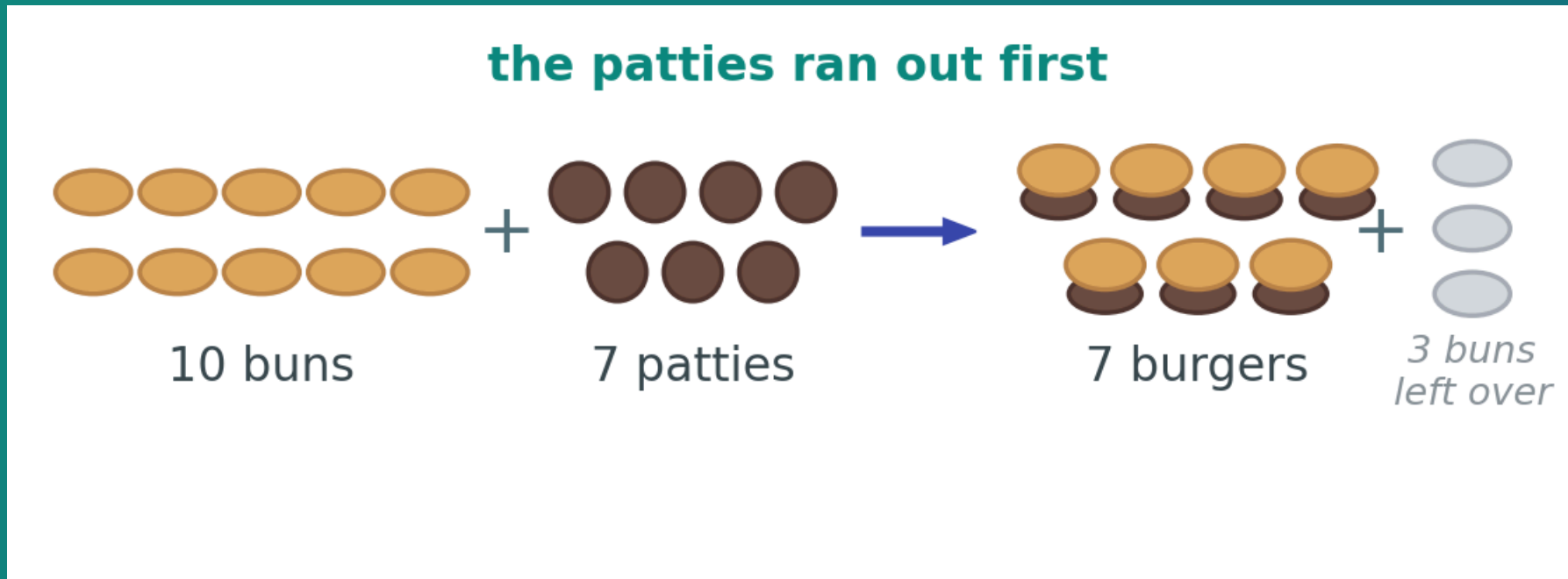
A stopped at moles: $5.00/114.22 \times (16/2) = 0.350 \text{ mol } \text{CO}_2$, one factor short of a count. C flipped Avogadro's number: $0.350 \div 6.022 \times 10^{23} = 5.82 \times 10^{-25}$, less than one molecule. D skipped the mole ratio: $5.00/114.22 \times 6.022 \times 10^{23} = 2.64 \times 10^{22}$.

 **Sense check:** Eight CO_2 per octane turns 0.0438 mol of fuel into 0.350 mol of CO_2 , about a third of a mole, and a third of 6.022×10^{23} is about 2×10^{23} ✓

3 · Limiting Reactant

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

One ingredient runs out first

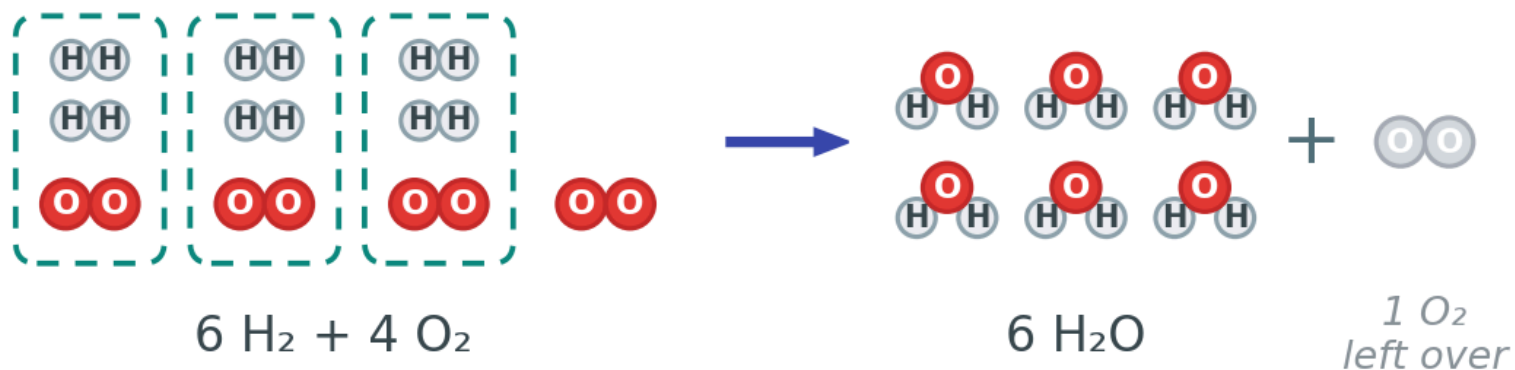


One bun and one patty per burger. The patties ran out first, so only seven burgers can be made. Reactions work the same way, in moles.

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.

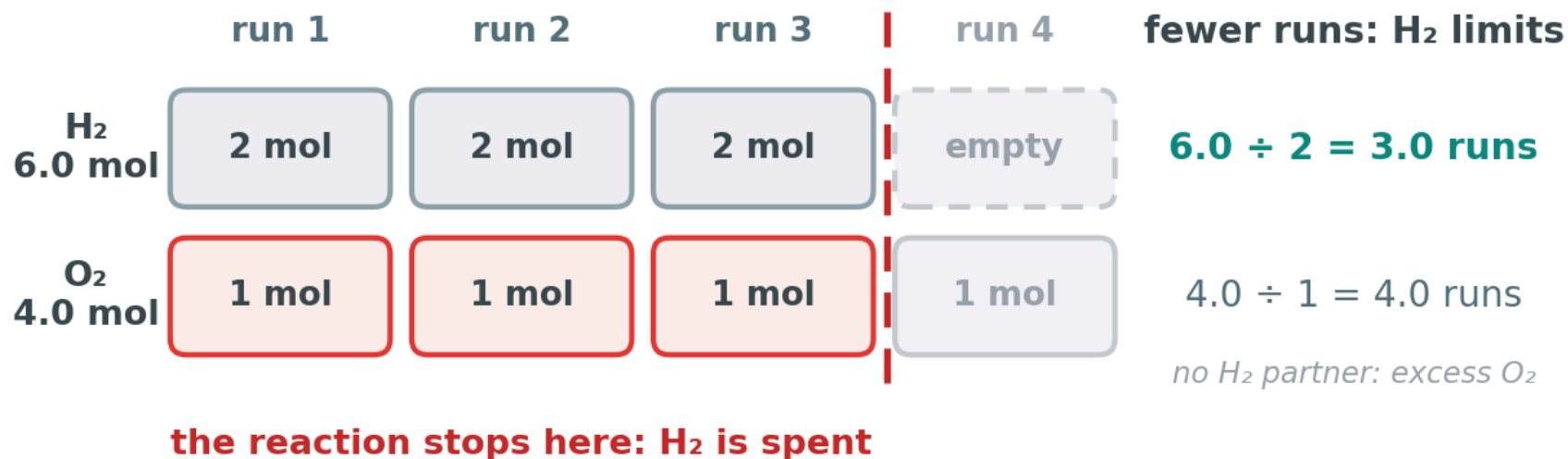
The test: moles $\div$ coefficient



H_2 : $6.0 \div 2 = 3.0$ runs · O_2 : $4.0 \div 1 = 4.0$ runs · $3.0 < 4.0$, 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





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

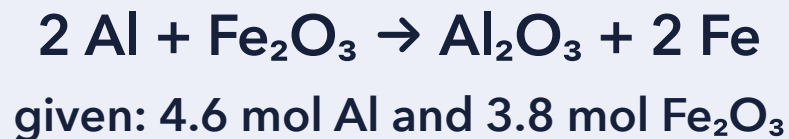


given: 4.6 mol Al and 3.8 mol Fe_2O_3

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 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

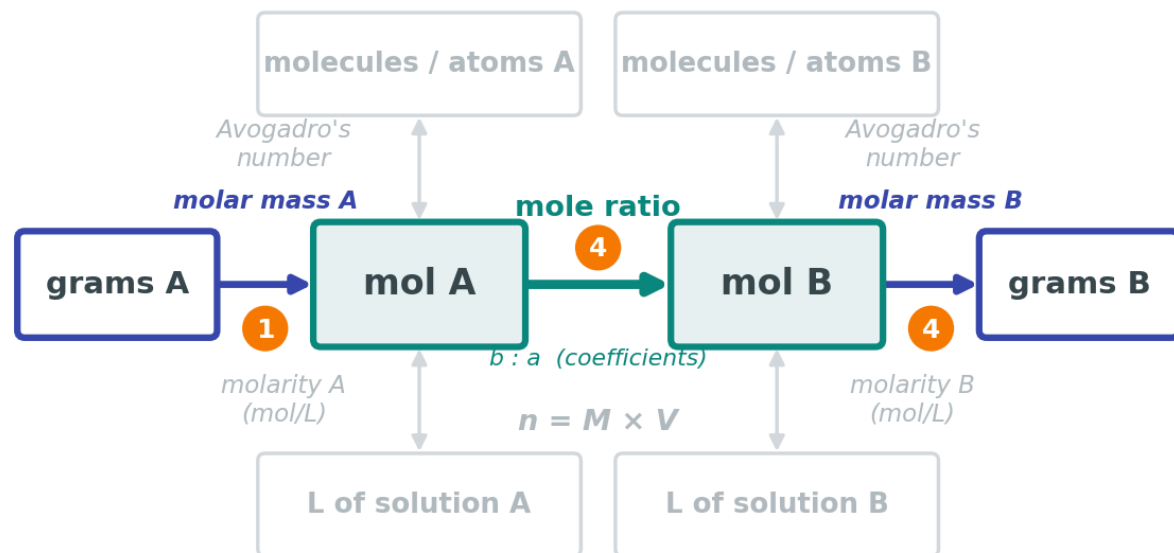
$$1.50 \text{ mol N}_2 \times \frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2} \times \frac{17.03 \text{ g NH}_3}{1 \text{ mol NH}_3} = 51.1 \text{ g NH}_3$$

 **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



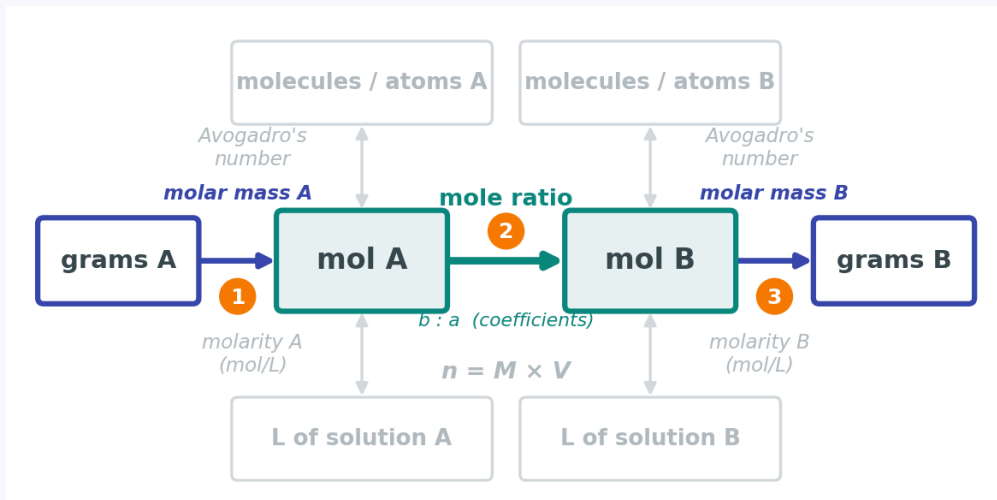
given: 42.0 g N₂ and 12.0 g H₂ · found: 51.1 g NH₃



🧠 **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. ✓



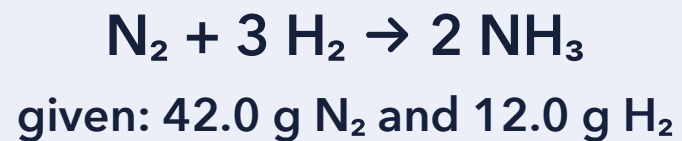
A second path: compare the product



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**.



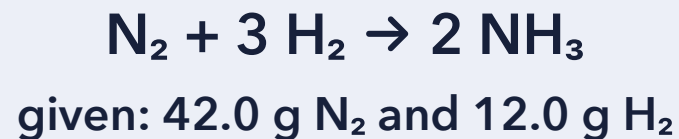
Worked example 3: the product path



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



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 NH_3 , then compare



given: 42.0 g N_2 and 12.0 g H_2 · $\text{N}_2 \rightarrow 3.00 \text{ mol NH}_3$ · $\text{H}_2 \rightarrow 3.97 \text{ mol NH}_3$

Step 3 • Moles to grams of product

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

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

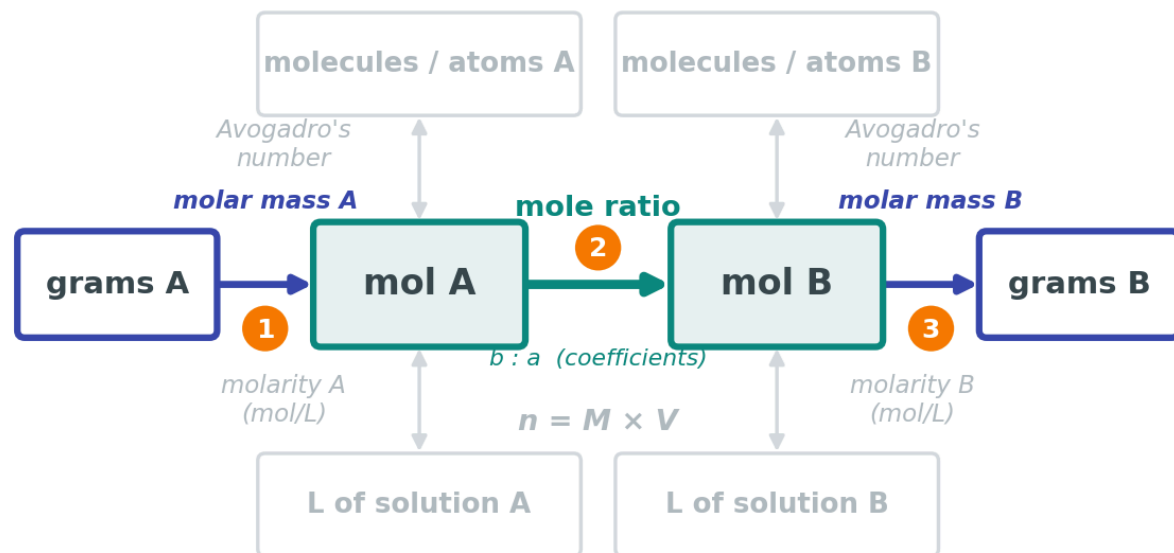
Step 4 • The reactant that makes less product limits

🧠 **Sense check:** 51.1 g < 67.6 g, so **N_2 limits**. The theoretical yield is **51.1 g NH_3** ; the 67.6 g from H_2 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 ✓

✓ Worked example 3: needed vs loaded



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

A third check: amount needed vs amount loaded

Convert one reactant into the amount of the other that it needs:

$$1.50 \text{ mol N}_2 \times \frac{3 \text{ mol H}_2}{1 \text{ mol N}_2} = 4.50 \text{ mol H}_2 \text{ needed}$$

🧠 **Sense check:** 4.50 mol H₂ needed, 5.95 mol loaded: enough H₂. N₂ runs out first, so **N₂ limits**, the same verdict. Needed more than loaded would mean H₂ limits ✓

Your turn: finish the comparison



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

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

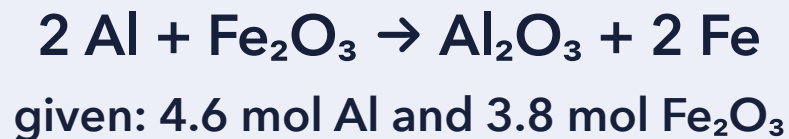
Limiting reactant:

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

⚠ 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 g CH₃OH (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

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

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

🧠 **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

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

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

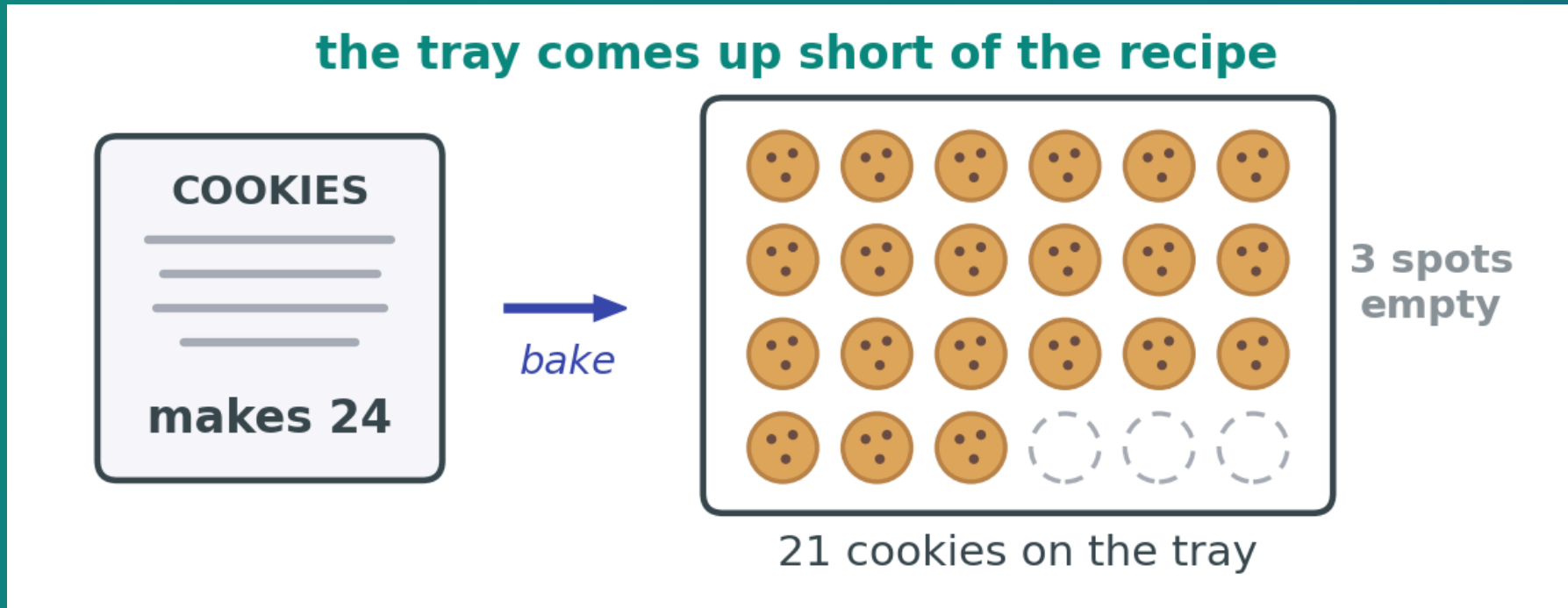
A stopped at the Mg consumed, 12.82 g. B inverted the ratio: $20.00 - 50.00/189.67 \times (1/2) \times 24.31 = 16.80$. C skipped it: $20.00 - 50.00/189.67 \times 24.31 = 13.59$.

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

4 • 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

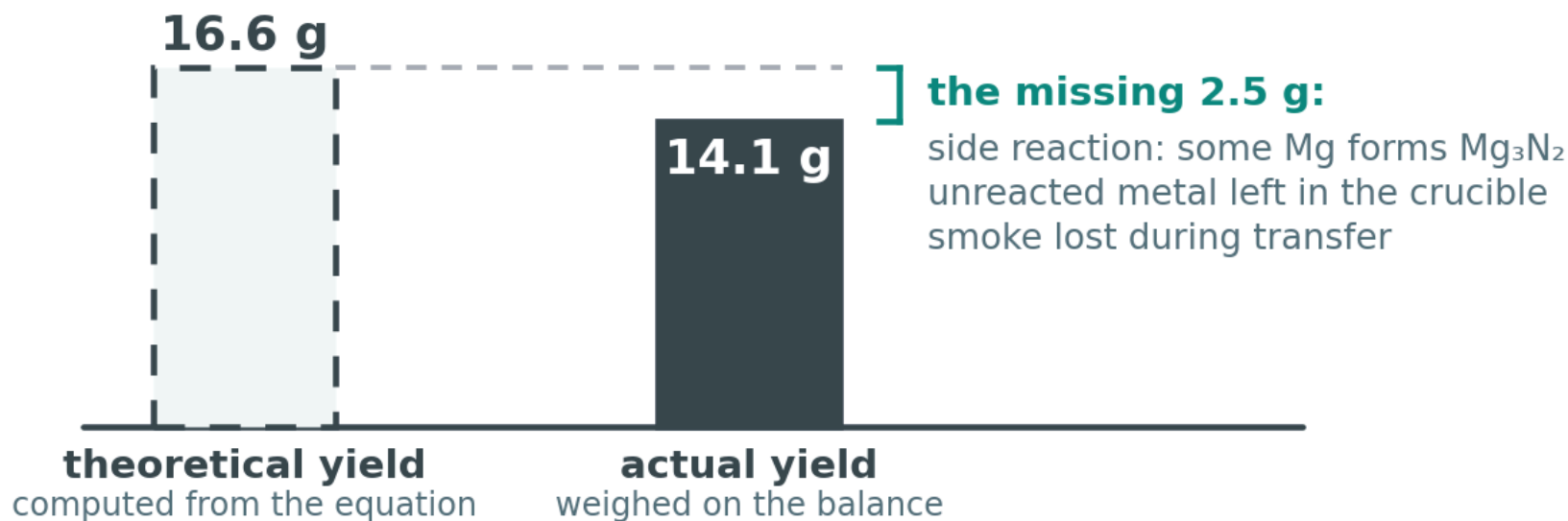


theoretical: 16.6 g MgO (computed) · actual: 14.1 g MgO (weighed)

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 Fe₂O₃, 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%.

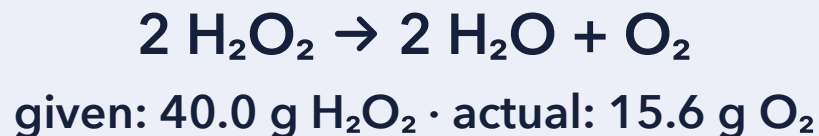
Reporting the fraction as the percent. $18.5 \div 22.0 = 0.841$ is a fraction of the maximum. Multiplied by 100 it becomes the percent yield, 84.1%. An answer of 0.841% would mean nearly everything was lost.

 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



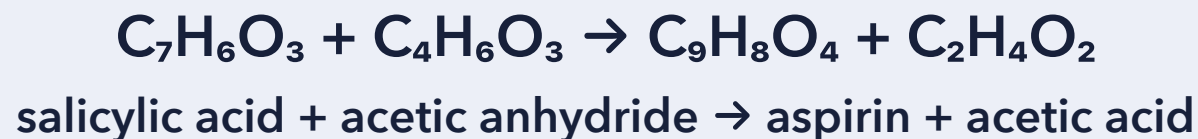
$$40.0 \text{ g } \text{H}_2\text{O}_2 \times \frac{1 \text{ mol } \text{H}_2\text{O}_2}{34.02 \text{ g } \text{H}_2\text{O}_2} \times \frac{1 \text{ mol } \text{O}_2}{2 \text{ mol } \text{H}_2\text{O}_2} \times \frac{32.00 \text{ g } \text{O}_2}{1 \text{ mol } \text{O}_2} = 18.81 \text{ g } \text{O}_2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

$$\frac{15.6 \text{ g } \text{O}_2}{18.81 \text{ g } \text{O}_2} \times 100 = \mathbf{82.9\% \text{ (answer B)}}$$

 **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 ✓

Check yourself

1. Percent yield divides two masses. Which is computed, and which is read from the balance? (*Which substance do both refer to?*)
2. Masses of **two** reactants are given, plus the product mass collected. List the steps from the given data to the percent yield. (*Which reactant sets the 100% mark?*)

Many of these reactions run in water, dosed by volume from a labeled solution. Molarity converts liters poured to moles delivered, the mole ratio takes over from there, and the same chains predict the product: solution stoichiometry.

Can you...?

- ☐ read the coefficients of a balanced equation as mole ratios and convert moles of one substance to moles of another?
- ☐ convert a mass of one substance to the mass, moles, or number of particles of another through the mole ratio?
- ☐ identify the limiting reactant, the maximum product it allows, and the excess reactant left over?
- ☐ calculate percent yield from the theoretical and actual yields?

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