

Thermochemistry

Preparation for General Chemistry · Dr. Karmach

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

- Tell kinetic from potential energy and convert energy among joules, calories, kilojoules, and food Calories
- Classify a process as exothermic or endothermic from the sign of ΔH , a heat term in the equation, or what is observed
- Use $q = m \cdot c \cdot \Delta T$ to find heat, mass, specific heat, or temperature change, and compare how specific heats set the temperature response to the same heat
- Apply heat lost = heat gained in a coffee-cup calorimeter to find the heat change of a metal, its specific heat, or the heat of a reaction in solution
- Find the energy content of a food from the water's temperature rise in a bomb calorimeter, in kJ and Calories per gram
- Write a thermochemical equation and use ΔH as a conversion factor for moles or grams of any substance in it

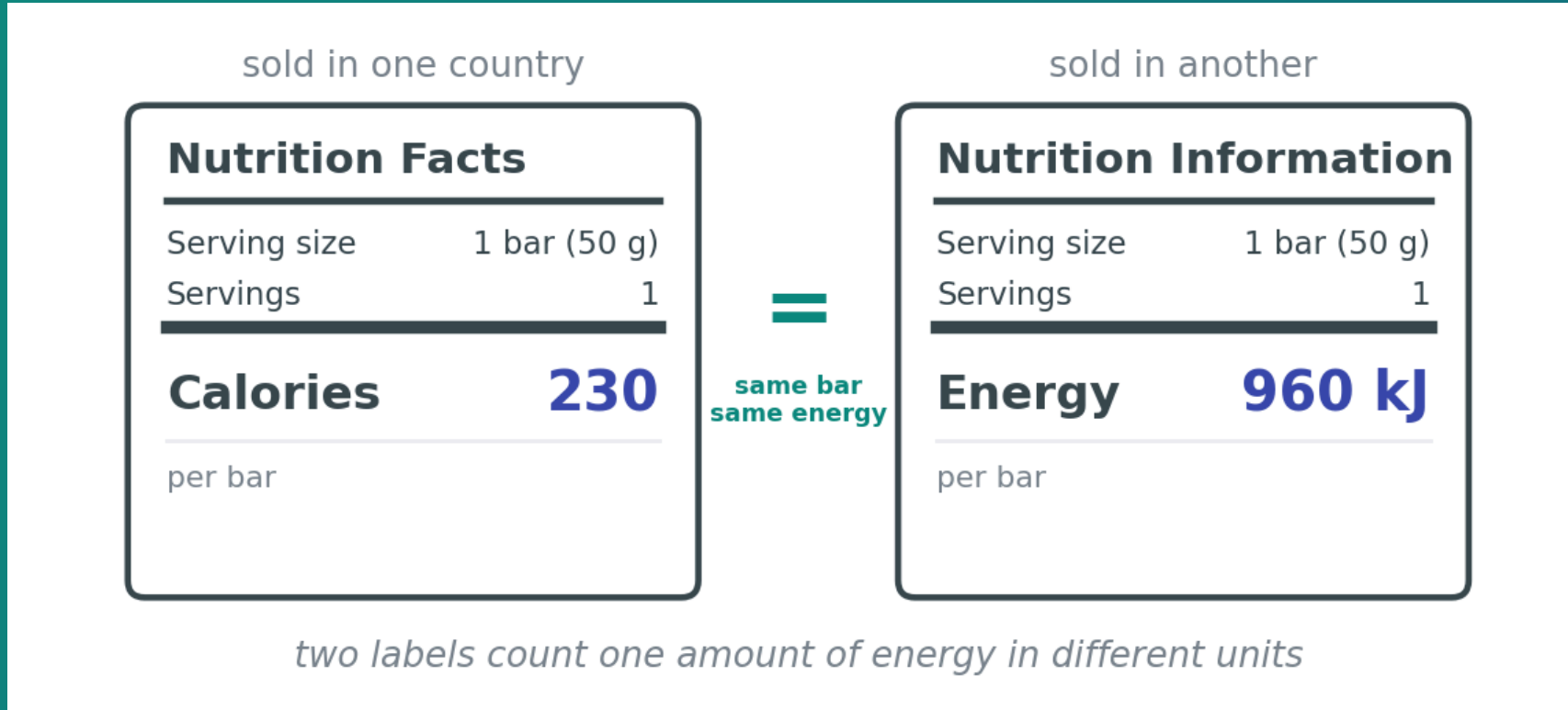
Today's route

1. **Energy & Its Units**
2. **Exothermic & Endothermic**
3. **Specific Heat & $q = mc\Delta T$**
4. **Comparing Specific Heats**
5. **Calorimetry**
6. **Thermochemical Equations**

1 • Energy & Its Units

Tell kinetic from potential energy and heat from temperature, and convert any amount of energy among joules, calories, kilojoules, and food Calories.

Same snack, two labels



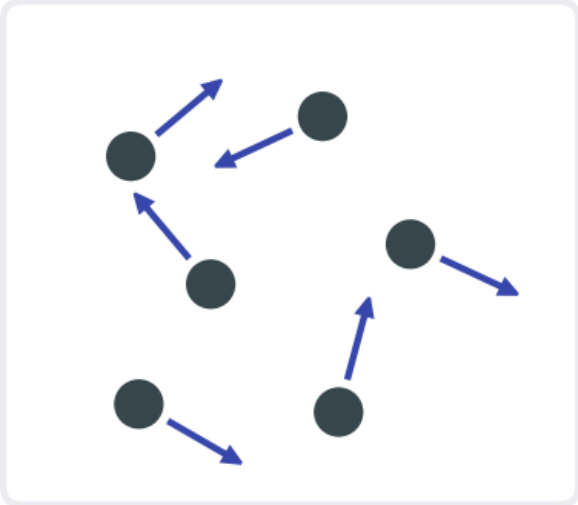
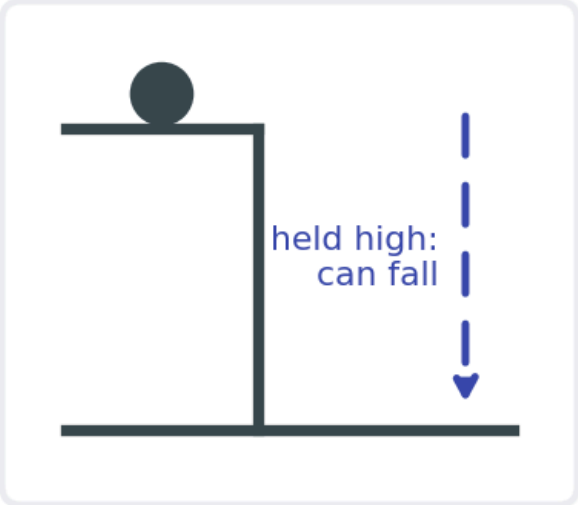
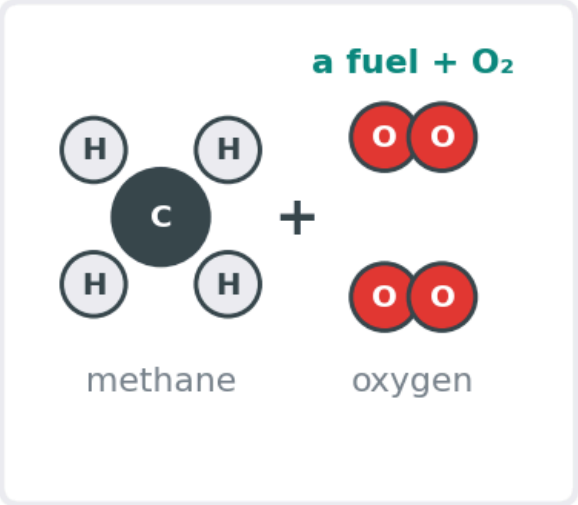
The same snack bar sells in two countries. One label lists **230 Calories**; the other lists **960 kJ**. Both describe the same energy, counted in different units.

Energy is never created or destroyed

Energy is the capacity to transfer heat or to do work. A process moves energy from one place to another or changes its form. The total amount never changes.

energy leaving one place = energy arriving in another
the total is fixed, so an amount of energy can be counted, like mass

Kinetic and potential energy

kinetic energy	potential energy	chemical potential energy
		
energy of motion	stored by position	stored in the arrangement of atoms

Kinetic energy is energy of motion. **Potential energy** is energy stored by position or arrangement. Fuels hold **chemical potential energy**: stored in the arrangement of atoms, released when a reaction rearranges them.

Two ways energy transfers: heat and work

Energy transfers between things in exactly two ways. **Heat (q)** is transfer driven by a temperature difference. **Work (w)** is transfer by a force moving something.

heat (q): hot pan $\rightarrow$ cool water
energy flows because the temperatures differ

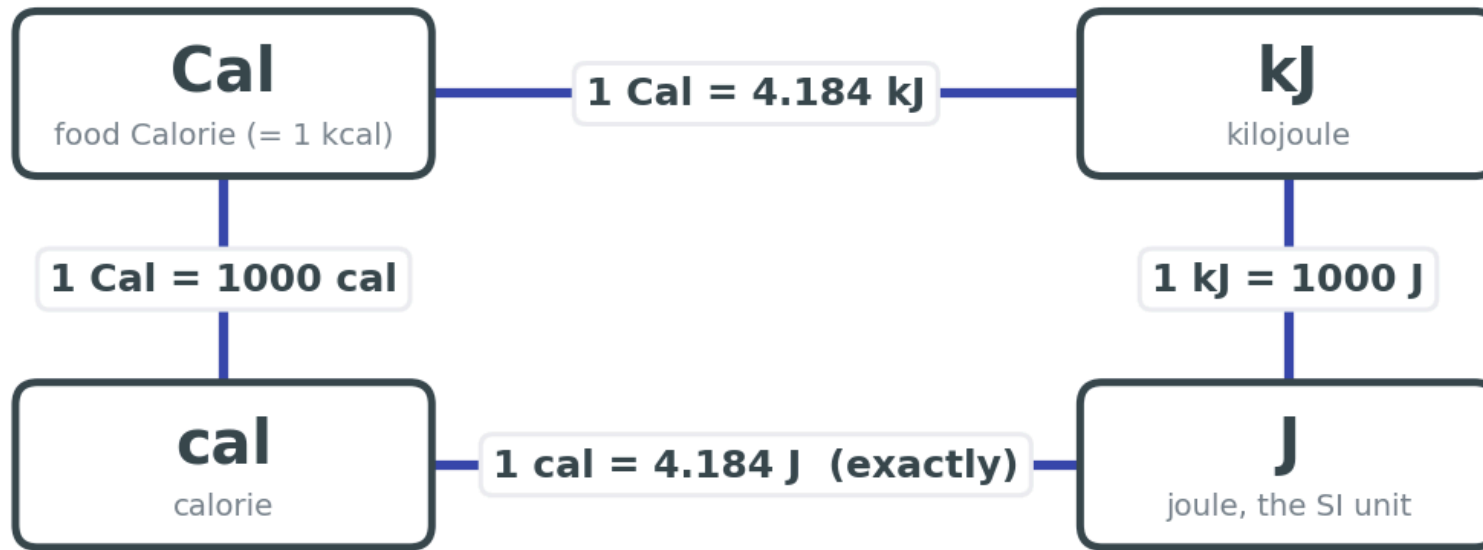
work (w): expanding gas pushes a piston
energy moves because a force acts through a distance

Temperature is not heat

Temperature measures the average kinetic energy of the particles: an intensity. **Heat** is an amount of energy in transfer. More sample means more energy at the same temperature.

a cup and a bathtub, both at 40 °C
same temperature, yet the tub transfers far more heat as it cools

The units of energy



each edge is an equality, and each equality gives two conversion factors

The SI unit is the **joule (J)**. $1 \text{ cal} = 4.184 \text{ J}$, exactly. The food **Calorie** has a capital C: $1 \text{ Cal} = 1 \text{ kcal} = 1000 \text{ cal}$. Each equality is an ordinary conversion factor.



The method

1. **Write the given:** number and unit.
2. **Plan the route:** given unit $\rightarrow$ wanted unit. Read a capital C as 1000 cal.
3. **Chain the factors so each unit cancels.**
4. **Sense-check** the size and the surviving unit.



Worked example 1: calories to joules

Step 1 • Write the given

$$1 \text{ cal} = 4.184 \text{ J}$$

given: 175 cal • wanted: J

A single-use hand warmer releases **175 cal** of heat as the iron inside it rusts. Express the energy in joules.

✓ Worked example 1: solution

$$1 \text{ cal} = 4.184 \text{ J}$$

given: 175 cal · wanted: J

Step 2 · Plan the route

Step 3 · Chain the factors so each unit cancels

$$175 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = 732 \text{ J}$$

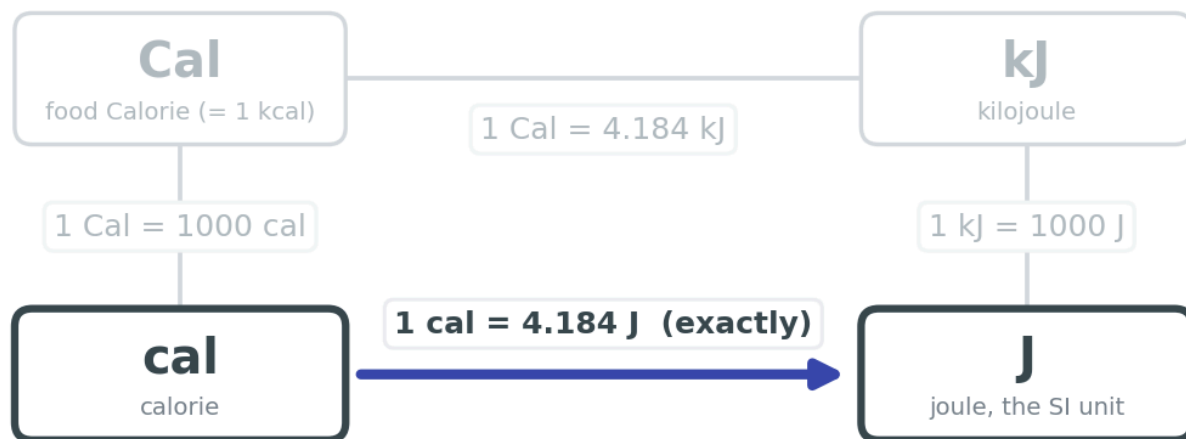
Step 4 · Sense-check

 **Sense check:** A joule is smaller than a calorie, so the count in joules must be larger: 175 → 732. The heat itself is unchanged. ✓

✓ Worked example 1: the route on the map

$$1 \text{ cal} = 4.184 \text{ J}$$

given: 175 cal · found: 732 J



route: cal → J · one conversion factor

🧠 **Sense check:** The units cal and J share one edge. One equality gives one conversion factor. ✓



Worked example 2: kilojoules to calories

Step 1 • Write the given

$$1 \text{ kJ} = 1000 \text{ J} \cdot 1 \text{ cal} = 4.184 \text{ J}$$

given: 2.50 kJ • wanted: cal

An instant cold pack absorbs **2.50 kJ** of heat from the skin it touches. Express the energy in calories.

✓ Worked example 2: solution

$$1 \text{ kJ} = 1000 \text{ J} \cdot 1 \text{ cal} = 4.184 \text{ J}$$

given: 2.50 kJ · wanted: cal · plan: kJ → J → cal

Step 2 · Plan the route

Step 3 · Chain the factors so each unit cancels

$$2.50 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 598 \text{ cal}$$

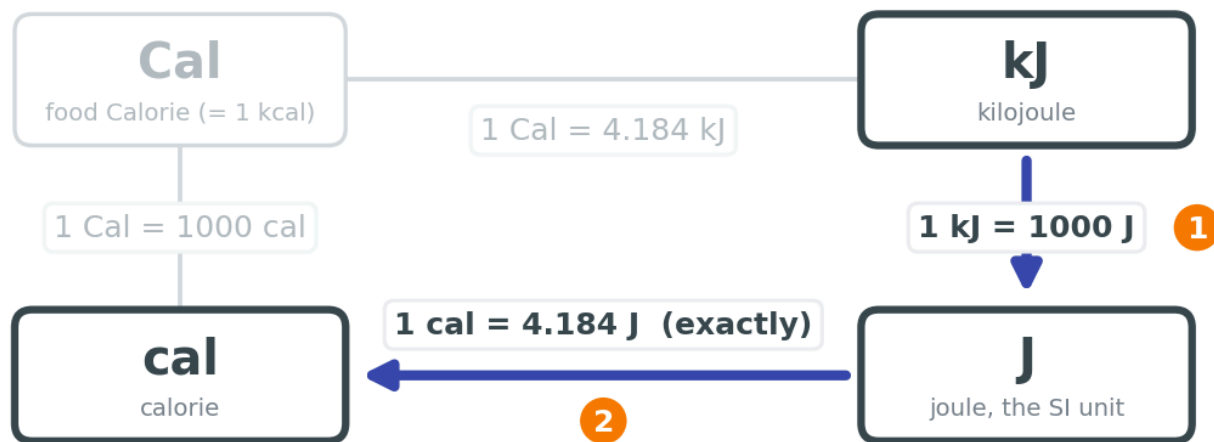
Step 4 · Sense-check

 **Sense check:** One calorie holds 4.184 J, so 2500 J make fewer calories than joules: 598. The pack absorbs the same heat under either name. ✓

✓ Worked example 2: the route on the map

$$1 \text{ kJ} = 1000 \text{ J} \cdot 1 \text{ cal} = 4.184 \text{ J}$$

given: 2.50 kJ · found: 598 cal



route: kJ → J → cal · two conversion factors

🧠 **Sense check:** No edge joins kJ to cal. The route takes two edges through the joule: two conversion factors. ✓

Your turn: calories to kilojoules

$$1 \text{ cal} = 4.184 \text{ J} \cdot 1 \text{ kJ} = 1000 \text{ J}$$

given: $7.10 \times 10^3 \text{ cal}$ · wanted: kJ · plan: cal → J → kJ

Burning one gram of ethanol releases $7.10 \times 10^3 \text{ cal}$.

$$7.10 \times 10^3 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{\text{kJ}}{\text{J}} = \text{kJ}$$

Fill the second factor from $1 \text{ kJ} = 1000 \text{ J}$, then compute.

$$7.10 \times 10^3 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 29.7 \text{ kJ}$$

⚠ Where this goes wrong

Writing the 4.184 factor upside down. $175 \text{ cal} \times (1 \text{ cal} / 4.184 \text{ J}) = 41.8 \text{ cal}^2/\text{J}$. No unit cancels, and the answer is not in joules. The factor that cancels cal gives 732 J.

Stopping at joules. The plan $\text{cal} \rightarrow \text{J} \rightarrow \text{kJ}$ has two arrows. Stopping after one leaves $2.97 \times 10^4 \text{ J}$: joules, not the wanted kilojoules. The chain ends at 29.7 kJ.

Reading temperature as an amount of energy. A cup of tea and a bathtub of water can both read 40°C . The temperatures match; the tub holds far more energy. Temperature is an intensity; heat is an amount.



Practice 1

$$1 \text{ cal} = 4.184 \text{ J}$$

given: $1.10 \times 10^3 \text{ J}$ · wanted: cal

A burning kitchen match releases $1.10 \times 10^3 \text{ J}$ of heat. How many calories is that?

- A. 1.10×10^3
- B. 0.263
- C. 263
- D. 4.60×10^3

✓ Practice 1 · answer: C

$$1 \text{ cal} = 4.184 \text{ J}$$

given: $1.10 \times 10^3 \text{ J}$ · plan: $\text{J} \rightarrow \text{cal}$

$$1.10 \times 10^3 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 263 \text{ cal (answer C)}$$

D flipped the 4.184 factor: $1.10 \times 10^3 \times 4.184 = 4.60 \times 10^3$, and no unit cancels. A skipped the conversion: 1.10×10^3 is still the count of joules. B also divided by 1000, as if the wanted unit were the food Calorie: $1.10 \times 10^3 / 4.184 / 1000 = 0.263$.

 **Sense check:** Each calorie holds 4.184 J, so the count in calories is smaller than the count in joules: $1.10 \times 10^3 \rightarrow 263$. ✓



Practice 2

$$1 \text{ kJ} = 1000 \text{ J} \cdot 1 \text{ cal} = 4.184 \text{ J}$$

given: 6.20 kJ · wanted: cal

Dissolving calcium chloride in a beaker of water releases **6.20 kJ** of heat. How many calories is that?

- A. 2.59×10^4
- B. 1.48×10^3
- C. 6.20×10^3
- D. 1.48

✓ Practice 2 · answer: B

$$1 \text{ kJ} = 1000 \text{ J} \cdot 1 \text{ cal} = 4.184 \text{ J}$$

given: 6.20 kJ · plan: kJ → J → cal

$$6.20 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 1.48 \times 10^3 \text{ cal (answer B)}$$

A flipped the 4.184 factor: $6.20 \times 10^3 \times 4.184 = 2.59 \times 10^4$, and no unit cancels. C stopped after the prefix factor: $6.20 \times 1000 = 6.20 \times 10^3$, a count of joules, not calories. D treated kilojoules as joules: $6.20 / 4.184 = 1.48$, a thousand times too small.

 **Sense check:** Each calorie holds 4.184 J, so $6.20 \times 10^3 \text{ J}$ make fewer calories than joules: 1.48×10^3 . ✓

Worked example 3: the food Calorie

Step 1 • Write the given

$$1 \text{ Cal} = 1000 \text{ cal} \cdot 1 \text{ cal} = 4.184 \text{ J} \cdot 1 \text{ kJ} = 1000 \text{ J}$$

given: 230 Cal • wanted: kJ

A snack bar's label lists **230 Calories**; the same bar abroad is labeled **960 kJ**. Express 230 Cal in kilojoules.

A common first attempt treats 230 Calories as 230 calories. Test it.

✓ Worked example 3: solution

$$1 \text{ Cal} = 1000 \text{ cal} \cdot 1 \text{ cal} = 4.184 \text{ J} \cdot 1 \text{ kJ} = 1000 \text{ J}$$

given: 230 Cal · wanted: kJ

A common first attempt

$$230 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 0.962 \text{ kJ} \times$$

🧠 **Sense check:** The kilojoule label reads 960: this result is 1000 times too small. The label's unit is Cal, not cal. ✗

✓ Worked example 3: the correct chain

$$1 \text{ Cal} = 1000 \text{ cal} \cdot 1 \text{ cal} = 4.184 \text{ J} \cdot 1 \text{ kJ} = 1000 \text{ J}$$

given: 230 Cal · wanted: kJ

Step 2 · Plan the route

Step 3 · Chain the factors so each unit cancels

$$230 \text{ Cal} \times \frac{1000 \text{ cal}}{1 \text{ Cal}} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 962 \text{ kJ}$$

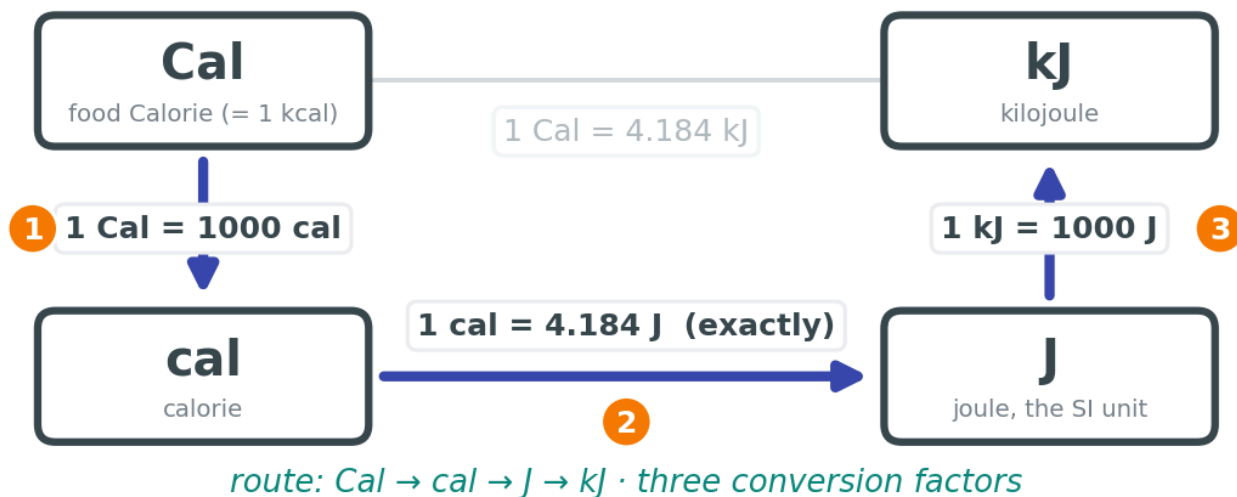
Step 4 · Sense-check

 **Sense check:** The kilojoule label reads 960: the chain reproduces it, rounded. The same chain shows 1 Cal = 4.184 kJ. ✓

✓ Worked example 3: the route on the map

$$1 \text{ Cal} = 1000 \text{ cal} \cdot 1 \text{ cal} = 4.184 \text{ J} \cdot 1 \text{ kJ} = 1000 \text{ J}$$

given: 230 Cal · found: 962 kJ



 **Sense check:** Three edges, three conversion factors. The trip around the map reproduces the top edge: 1 Cal = 4.184 kJ. ✓

Take-home: the food Calorie is a kilocalorie

Do: read the capital C as 1000 cal. Big C, big unit.

$$230 \text{ Cal} \times (1000 \text{ cal} / 1 \text{ Cal}) \times (4.184 \text{ J} / 1 \text{ cal}) \times (1 \text{ kJ} / 1000 \text{ J}) = 962 \text{ kJ}$$

matches the 960 kJ label ✓

Do not: read Cal as cal. The answer lands 1000 times too small.

$$230 \text{ cal} \times (4.184 \text{ J} / 1 \text{ cal}) \times (1 \text{ kJ} / 1000 \text{ J}) = 0.962 \text{ kJ}$$

1000 times smaller than the label ✗



Practice 3

$$1 \text{ Cal} = 1000 \text{ cal} \cdot 1 \text{ cal} = 4.184 \text{ J} \cdot 1 \text{ kJ} = 1000 \text{ J}$$

given: 205 Cal · wanted: cal, J and kJ

One cup of cooked rice provides **205 Calories**. How much energy is that in calories, in joules and in kilojoules?

- A. cal: 205 · J: 858 · kJ: 0.858
- B. cal: 2.05×10^5 · J: 8.58×10^5 · kJ: 858
- C. cal: 2.05×10^5 · J: 4.90×10^4 · kJ: 49.0
- D. cal: 2.05×10^5 · J: 858 · kJ: 858
- E. cal: 2.05×10^5 · J: 8.58×10^5 · kJ: 8.58×10^8

✓ Practice 3 · answer: B

$$205 \text{ Cal} = 2.05 \times 10^5 \text{ cal} = 8.58 \times 10^5 \text{ J} = 858 \text{ kJ (answer B)}$$

given: $205 \text{ Cal} \cdot 1 \text{ Cal} = 1000 \text{ cal} \cdot 1 \text{ cal} = 4.184 \text{ J} \cdot 1 \text{ kJ} = 1000 \text{ J} \cdot \text{each hop reported}$

$$205 \text{ Cal} \times \frac{1000 \text{ cal}}{1 \text{ Cal}} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 858 \text{ kJ}$$

A read Cal as cal: each value a thousand times too small. C flipped 4.184: $2.05 \times 10^5 / 4.184 = 4.90 \times 10^4 \text{ J}$. D wrote the kJ number as the J count: $858 \text{ kJ} = 8.58 \times 10^5 \text{ J}$. E multiplied by 1000 at the last hop.

🧠 **Sense check:** After the chain, check with $1 \text{ Cal} = 4.184 \text{ kJ}$: $205 \times 4.184 = 858$. ✓



Practice 4

$$1 \text{ Cal} = 1000 \text{ cal} \cdot 1 \text{ cal} = 4.184 \text{ J} \cdot 1 \text{ kJ} = 1000 \text{ J}$$

given: 8.50 Cal per minute · 40.0 min · wanted: kJ

Swimming laps burns about **8.50 Calories per minute**. How many kilojoules do **40.0 minutes** of laps use?

- A. 1.42
- B. 81.3
- C. 340.
- D. 1.42×10^3

✓ Practice 4 • answer: D

$$1 \text{ Cal} = 1000 \text{ cal} \cdot 1 \text{ cal} = 4.184 \text{ J} \cdot 1 \text{ kJ} = 1000 \text{ J}$$

given: 8.50 Cal per minute · 40.0 min · plan: min → Cal → cal → J → kJ

$$40.0 \text{ min} \times \frac{8.50 \text{ Cal}}{1 \text{ min}} = 340. \text{ Cal}$$

$$340. \text{ Cal} \times \frac{1000 \text{ cal}}{1 \text{ Cal}} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 1.42 \times 10^3 \text{ kJ (answer D)}$$

A read Cal as cal: $340. \times 4.184 / 1000 = 1.42$, a thousand times too small. B flipped the 4.184 factor: $340. \times 1000 / 4.184 / 1000 = 81.3$, and the joule never cancels. C stopped at Calories: 340. is the energy in Cal, never converted to kJ.

 **Sense check:** One food Calorie is 4.184 kJ, so 340 Cal sits near $340 \times 4 = 1360 \text{ kJ}$. The chain gives 1.42×10^3 . ✓



Check yourself

1. A label lists **95 Calories**. Write the chain to kilojoules: which factor comes first, and what does each unit cancel into?
2. A cup of water and a pot of water both read 60 °C. Which quantity matches, and which differs: temperature, or energy content?

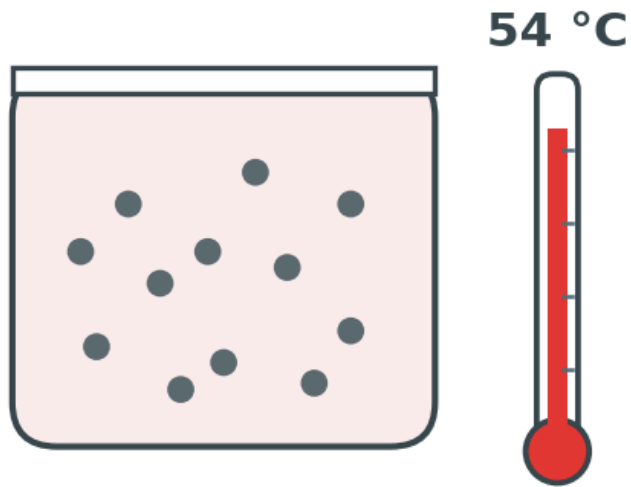
Every joule a system gains or loses arrives as heat (q) or as work (w). The first law adds them: $\Delta E = q + w$. At constant pressure the heat term earns its own name, the enthalpy change ΔH .

2 · Exothermic & Endothermic

Classify any process as exothermic or endothermic (from the sign of ΔH , from an energy diagram, or from a heat term written into the equation) and state which way heat flows between system and surroundings.

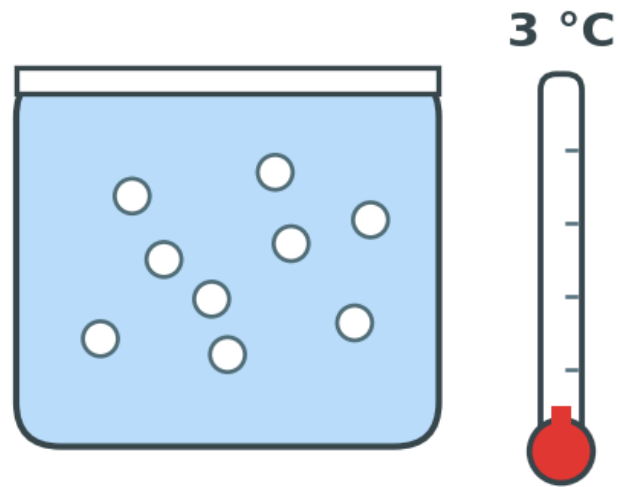
Two pouches from the drugstore

snap this pouch: it heats up



sealed pouch: iron powder + air

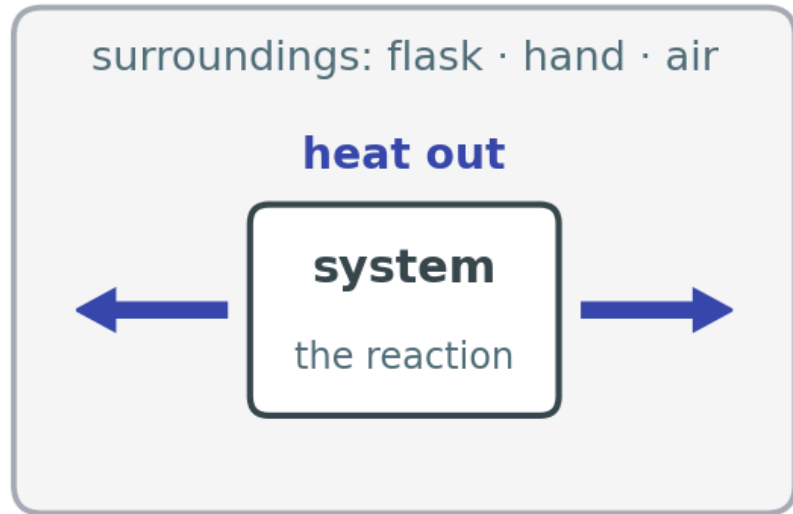
snap this pouch: it turns icy



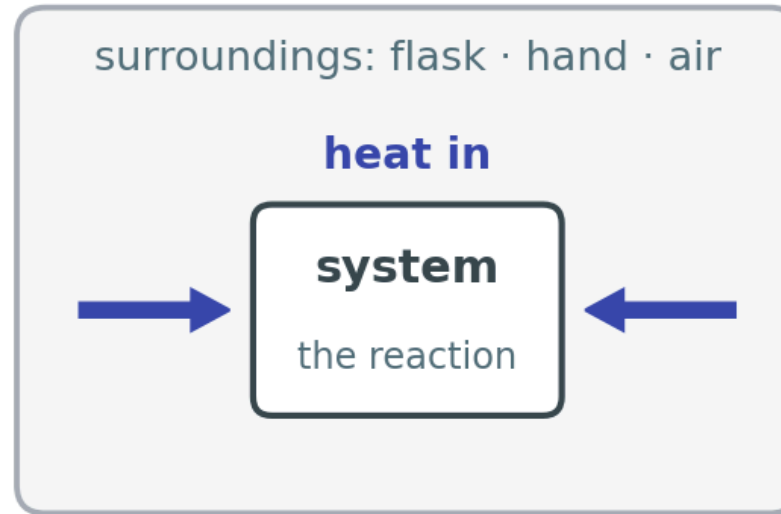
sealed pouch: water + salt pellets

Snap the pouch inside a hand warmer and it climbs to 54 °C. Snap a cold pack and it drops near freezing. Sealed chemicals drive both changes.

The system and its surroundings



heat out → ΔH negative
surroundings warm up



heat in → ΔH positive
surroundings cool down

The reaction is the **system**; the flask, your hand, the room are the **surroundings**. ΔH records the system's heat: out negative, in positive. The sign follows the system, not your hand.

Enthalpy

Enthalpy, H , is the heat content of a system. A change in it, ΔH , equals the heat of the process at constant pressure. An open flask or a pouch in your hand qualifies.

ΔH = heat of the process at constant pressure
heat out of the system $\rightarrow \Delta H$ negative · heat in $\rightarrow \Delta H$ positive

Exothermic and endothermic

exothermic: heat exits the system

surroundings warm up · ΔH negative · burning fuel, the hand-warmer pouch

endothermic: heat enters the system

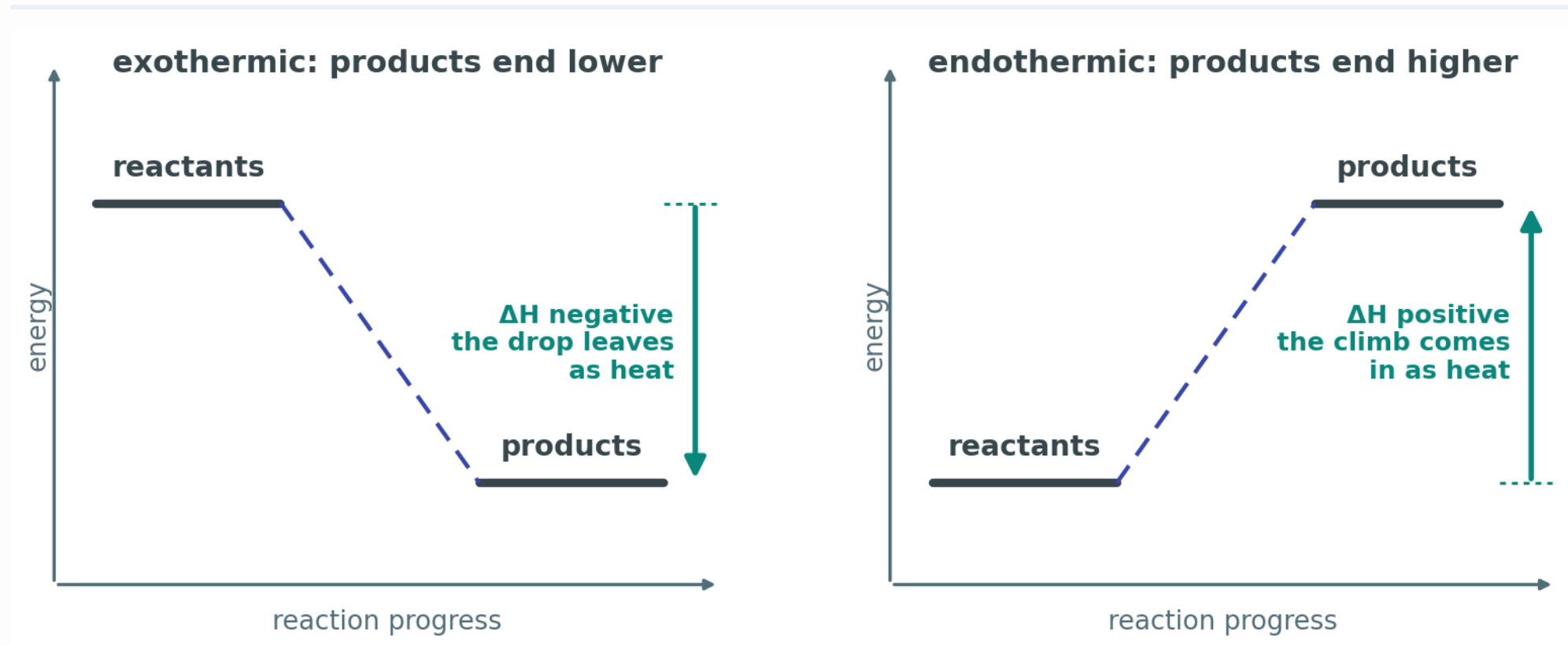
surroundings cool down · ΔH positive · melting ice, the cold-pack pouch

memory hook: EXo, heat EXits · ENdo, heat ENters

the prefix names the heat's direction, read from the system's side

Heat "flows from a hot object to a cold object". A process sending heat out is **exothermic**; one taking heat in is **endothermic**. Both names describe the system; the surroundings show the opposite change.

Energy diagrams



An energy diagram plots energy against reaction progress. Products below the reactants: the difference left as heat, exothermic. Products above: the difference came in as heat, endothermic.

Heat written into the equation



heat on the product side: it leaves with the products · $\Delta H = -890 \text{ kJ}$ · exothermic



heat on the reactant side: it must be supplied · $\Delta H = +572 \text{ kJ}$ · endothermic

A thermochemical equation may carry its heat in-line. Product side: heat released, exothermic. Reactant side: heat absorbed, endothermic. The separate ΔH states the same fact with a sign.



The method

1. **Name the system.** The process is the system; all else is surroundings.
2. **Find the heat's direction.** From the ΔH sign, diagram levels, or heat term.
3. **State the verdict.** Heat out: exothermic, ΔH negative. Heat in: endothermic, ΔH positive.



Worked example 1: a hand warmer's reaction



given: the equation and its ΔH · wanted: the verdict and the heat's direction

Inside a hand warmer, iron powder reacts with oxygen from the air.

Classify the reaction and state which way heat flows.

✓ Worked example 1: solution



given: the equation and its ΔH

Step 1 · Name the system

Step 2 · Find the heat's direction

Step 3 · State the verdict

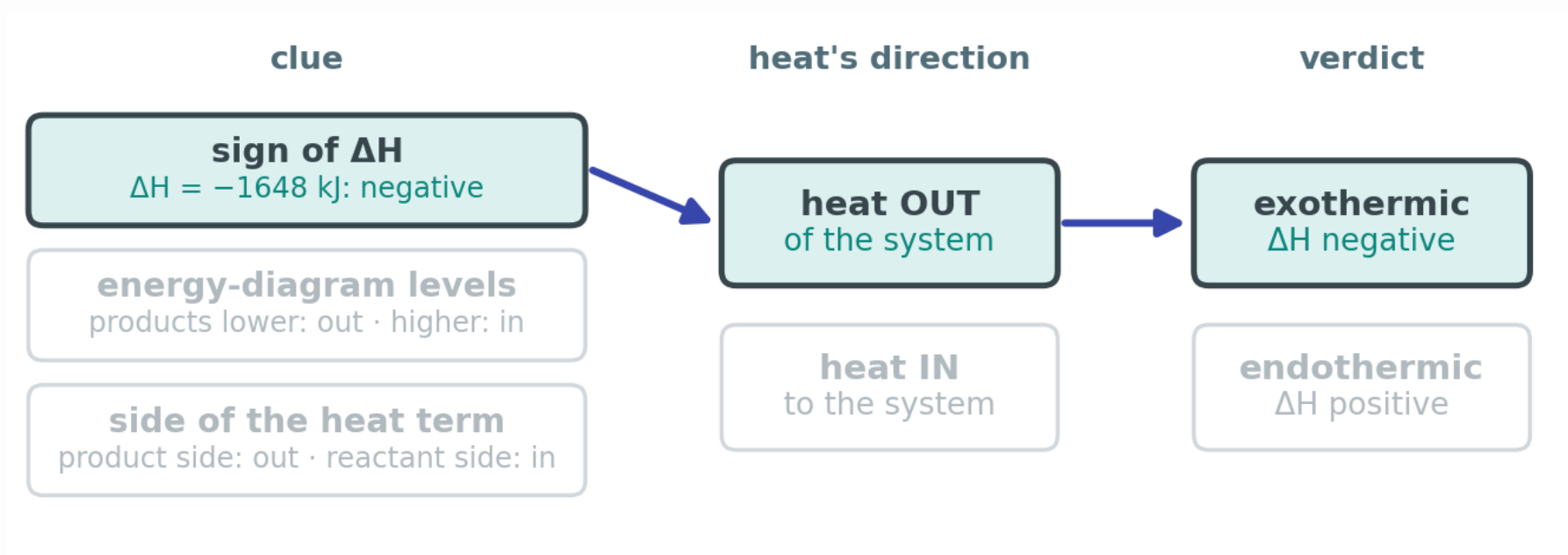
$4 \text{ Fe(s)} + 3 \text{ O}_2\text{(g)} \rightarrow 2 \text{ Fe}_2\text{O}_3\text{(s)} + 1648 \text{ kJ} \rightarrow \text{exothermic}$
 heat out · $\Delta H = -1648 \text{ kJ}$ · the surroundings (your hands) warm up

 **Sense check:** The pouch warms your hand: the surroundings gain exactly the heat the system loses. A negative ΔH and a warming hand agree.

✓ Worked example 1: the clue used



clue read: the sign of ΔH · verdict: exothermic

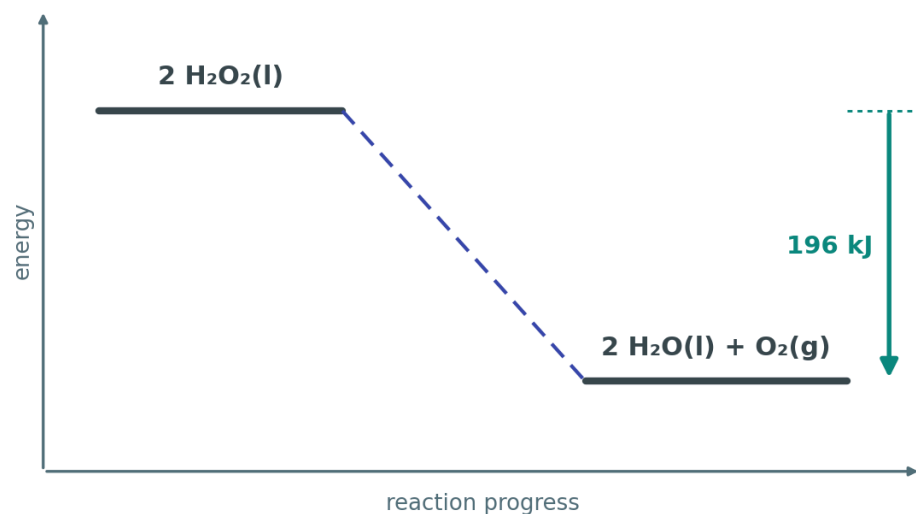


🧠 **Sense check:** Any one of the three clues gives the heat's direction. Here the sign settled it: negative, heat out, exothermic. ✓

Worked example 2: hydrogen peroxide decomposes

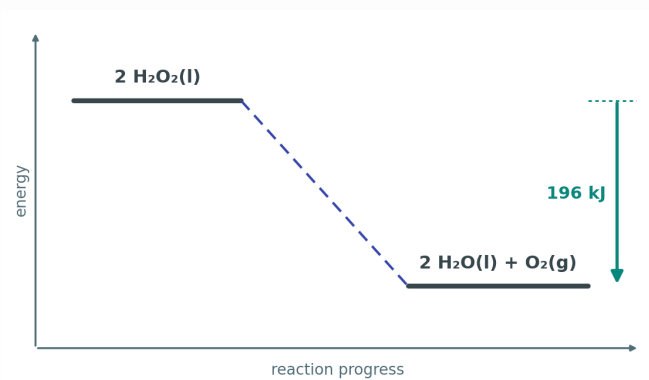


given: the energy diagram · wanted: the verdict and ΔH with its sign



Hydrogen peroxide fizzing on a cut breaks down into water and oxygen. Classify the process from the diagram and give ΔH .

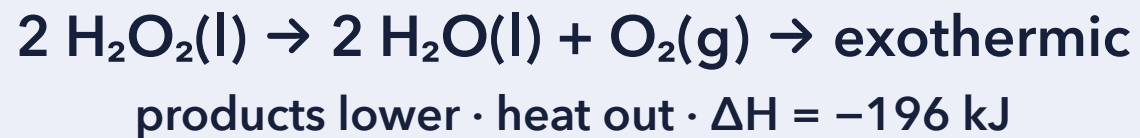
✓ Worked example 2: solution



Step 1 • Name the system

Step 2 • Find the heat's direction

Step 3 • State the verdict

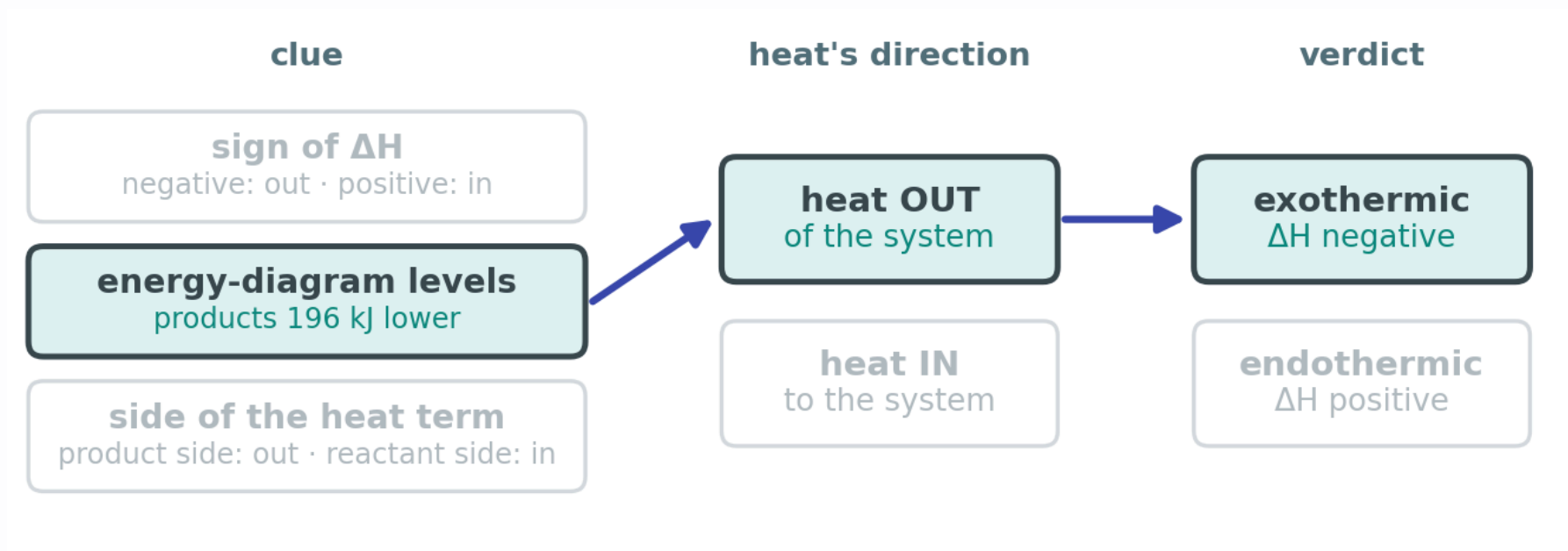


 **Sense check:** Downhill on an energy diagram is heat out. The products hold less energy than the reactants, and the fizzing cut warms slightly.

✓ Worked example 2: the clue used



clue read: the energy-diagram levels · verdict: exothermic, $\Delta H = -196 \text{ kJ}$



🧠 **Sense check:** No ΔH was written, and none was needed for the direction: products below the reactants is heat out.



Your turn: photosynthesis



a leaf builds glucose; the 2803 kJ arrives as sunlight

step	question	answer
1 · name the system	what is changing?	the CO_2 and water becoming glucose
2 · find the heat's direction	which side carries the heat term?	the _____ side; heat _____ the system
3 · state the verdict		_____ · $\Delta H =$ _____ kJ

Complete the three steps.

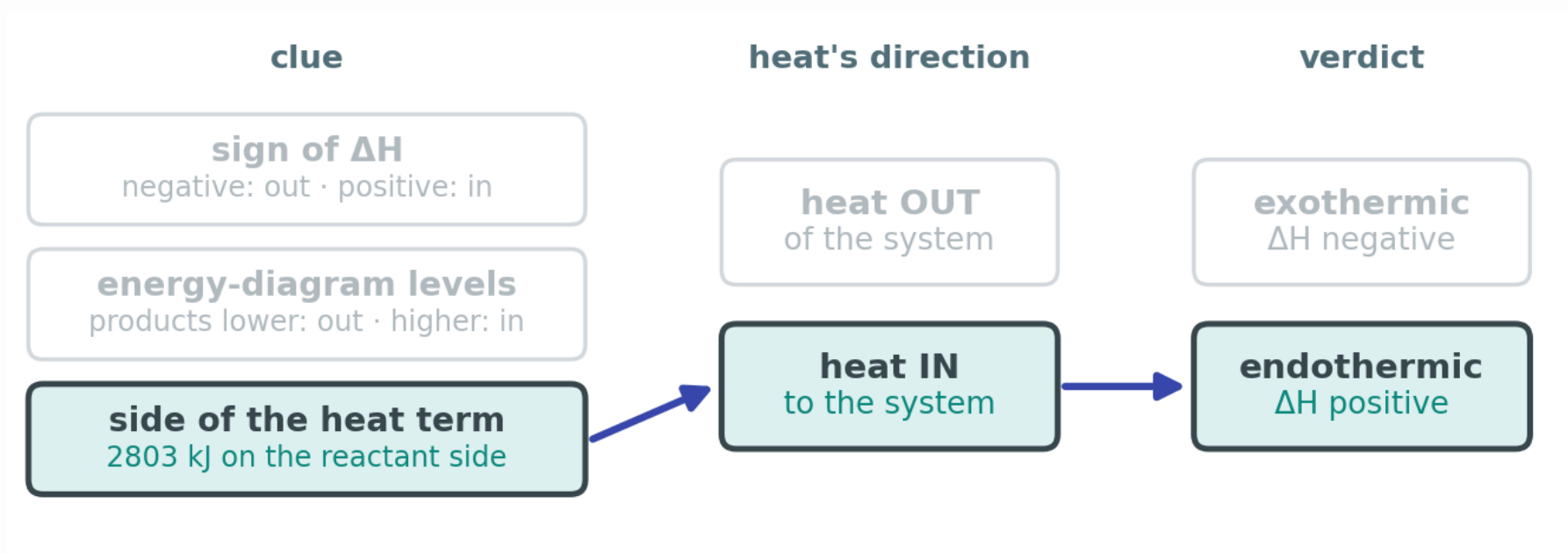
photosynthesis $\rightarrow$ endothermic

heat term on the reactant side · heat enters the system · $\Delta H = +2803 \text{ kJ}$

✓ Your turn: the clue used



clue read: the side of the heat term · verdict: endothermic, $\Delta H = +2803 \text{ kJ}$



 **Sense check:** A heat term with the reactants is heat the system must take in. Reactant side, heat in, endothermic. ✓

⚠ Where this goes wrong

Reading the sign from your hand. A cold pack chills your skin, and the chill gets recorded as heat lost: $\Delta H = -26 \text{ kJ}$. The skin is surroundings. Its loss is the system's gain: $\Delta H = +26 \text{ kJ}$.

Pairing a label with the opposite flow. "Exothermic, and it absorbs heat" contradicts itself. The label names the flow: exothermic releases heat, endothermic absorbs it. One verdict carries both parts.

Reading the heat term from the wrong side. In $2 \text{ H}_2\text{O}(\text{l}) + 572 \text{ kJ} \rightarrow 2 \text{ H}_2(\text{g}) + \text{O}_2(\text{g})$, the 572 kJ gets reported as released. It sits with the reactants, so it is consumed: absorbed, $\Delta H = +572 \text{ kJ}$.

Calling the higher level the bigger release. Height on an energy diagram is energy stored, not energy given off. Products above the reactants means the system took energy in: endothermic, ΔH positive.



Practice 1



given: the equation and its ΔH

Sulfur dioxide converts to sulfur trioxide during sulfuric acid manufacture. Which statement describes the reaction?

- A. Endothermic: heat is absorbed by the system from the surroundings
- B. Exothermic: heat is absorbed by the system from the surroundings
- C. Exothermic: heat is released by the system to the surroundings
- D. Endothermic: heat is released by the system to the surroundings

✓ Practice 1 · answer: C

$\Delta H = -198 \text{ kJ} \rightarrow \text{negative} \rightarrow \text{heat out} \rightarrow \text{exothermic (answer C)}$

heat released by the system · the surroundings warm up

A flipped the sign convention: heat absorbed would be counted into the system, +198 kJ, not –198 kJ. B paired the right label with the wrong flow: exothermic means heat exits the system. D paired the right flow with the wrong label: a heat-releasing reaction is exothermic.

 **Sense check:** Negative ΔH , heat out, exothermic, warmer surroundings: four readings of the same event.



Worked example 3: the cold pack



given: the pouch turns icy in your hand · wanted: the verdict and the sign of ΔH

Snapping the pack lets ammonium nitrate dissolve in water, and the pouch turns icy.

A common first answer: the pack is cold, so it is losing heat: exothermic. Test it.

✓ Worked example 3: solution



the pouch turns icy in your hand

A common first answer

cold pack, so the pack is losing heat → exothermic?

cold marks heat leaving the pack only if the pack were the surroundings ✕

Step 1 · Name the system

Step 2 · Find the heat's direction

Step 3 · State the verdict



heat in · measured $\Delta H = +26$ kJ per mole dissolved · the surroundings (your hand) cool

 **Sense check:** The pack feels cold *because* it absorbs heat. A cooling hand is the surroundings' loss and the system's gain: $\Delta H = +26$ kJ, never -26 kJ.

Take-home: your hand is the surroundings

feels hot → the surroundings are gaining heat → the system is losing it
exothermic · ΔH negative

feels cold → the surroundings are losing heat → the system is gaining it
endothermic · ΔH positive

Skin and thermometers sit in the surroundings. They report the surroundings' change, and the system did the opposite. Feels cold: the system is absorbing heat: endothermic, ΔH positive.

Practice 2: everyday changes

	process
1	water evaporates from a puddle
2	snow forms in a cloud
3	a strong acid is mixed with a strong base

Classify each process, in order, as exothermic or endothermic.

- A. endothermic · endothermic · exothermic
- B. endothermic · exothermic · exothermic
- C. exothermic · exothermic · exothermic
- D. endothermic · exothermic · endothermic

✓ Practice 2: answer B

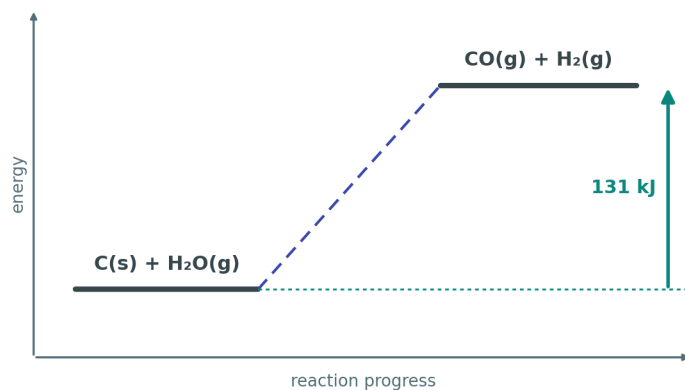
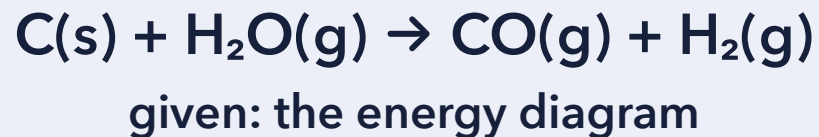
endothermic · exothermic · exothermic (answer B)

evaporating: heat in · water turning to ice: heat out · acid and base mixed: heat out

A took every change of state as heat in. Melting takes heat in, so its reverse, water turning to ice, gives heat out. C read evaporation's cooling as heat out. The cooling is the surroundings giving the evaporating water its heat. D assumed a reaction must absorb energy to happen; the acid and base mixture turns warm.

 **Sense check:** Evaporation cools the ground it leaves; an acid and base mixture warms its beaker. Each is the surroundings reporting, and the system does the opposite. ✓

Practice 3: water gas

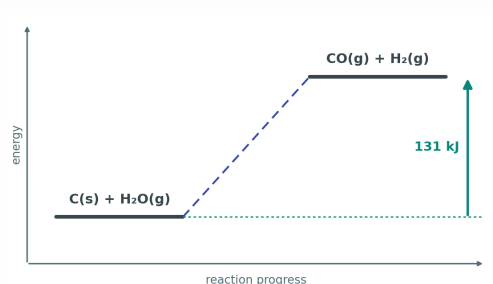


Steam over red-hot coke makes water gas, a fuel. Which statement describes the reaction?

- A. Exothermic, $\Delta H = -131 \text{ kJ}$
- B. Endothermic, $\Delta H = -131 \text{ kJ}$

- C. Exothermic, $\Delta H = +131 \text{ kJ}$
- D. Endothermic, $\Delta H = +131 \text{ kJ}$

✓ Practice 3: answer D



$\text{C(s)} + \text{H}_2\text{O(g)} \rightarrow \text{CO(g)} + \text{H}_2\text{(g)} \rightarrow$ endothermic, $\Delta H = +131 \text{ kJ}$ (answer D)
products 131 kJ above the reactants · the system took that energy in

A called the higher level the bigger release: height is energy stored, so products above the reactants took energy in. B took the sign from the surroundings: the furnace loses 131 kJ, and the system gains it, +131 kJ. C paired the positive sign with the wrong label: heat in is endothermic.

 **Sense check:** Uphill on an energy diagram is heat in. The coke must be kept red-hot for the reaction to run: endothermic, ΔH positive. ✓



Practice 4



given: the equation · the bucket turns hot

Water is stirred into quicklime and the bucket turns hot. A classmate writes: "The 65 kJ ends up with the products, so the system gains heat: endothermic, $\Delta H = +65 \text{ kJ}$." What is wrong with the reasoning, if anything?

- A. Heat written with the products is released: exothermic, $\Delta H = -65 \text{ kJ}$. The hot bucket is the surroundings.
- B. Nothing is wrong: heat written with the products is heat gained, so $\Delta H = +65 \text{ kJ}$.
- C. Only the label is wrong: $\Delta H = +65 \text{ kJ}$ stands, but a hot bucket means exothermic.
- D. Only the sign is wrong: it is endothermic, but heat that warms the bucket is -65 kJ .

✓ Practice 4 · answer: A

$\text{CaO(s)} + \text{H}_2\text{O(l)} \rightarrow \text{Ca(OH)}_2\text{(s)} + 65 \text{ kJ} \rightarrow \text{exothermic, } \Delta H = -65 \text{ kJ (answer A)}$
heat term on the product side: released · the hot bucket is the surroundings gaining that heat

B read a product-side heat term as heat gained: it leaves with the products, so it is released, -65 kJ , not $+65 \text{ kJ}$. C matched the sign but not the label: a positive ΔH would mean heat in, and a bucket cannot turn hot from that. D matched the label but not the sign: endothermic and $\Delta H = -65 \text{ kJ}$ contradict each other.

 **Sense check:** Two pieces of evidence, one verdict: the heat term sits with the products, and the surroundings warm. Heat out, exothermic, ΔH negative.



Extra practice 1



given: the flask rests on a wet wooden board · the water under it freezes

Two white solids are stirred together in a flask on a wet board, and within minutes the water under the flask freezes. Which statement describes the reaction?

- A. Exothermic, ΔH negative: heat flows from the flask into the board
- B. Exothermic, ΔH positive: heat flows from the board into the flask
- C. Endothermic, ΔH positive: heat flows from the board into the flask
- D. Endothermic, ΔH negative: heat flows from the board into the flask

✓ Extra practice 1 · answer: C



surroundings: the board and its water · heat flows board → flask · endothermic, ΔH positive (answer C)

A took the freezing as the reaction's own change. Freezing does release heat, but the water is surroundings, and its heat went into the flask. B paired the label with the wrong flow: heat into the system is endothermic. D read the sign from the surroundings: the board loses heat, the system gains it, so ΔH is positive.

 **Sense check:** Water freezes only when heat is pulled out of it. The flask pulled it out: the system absorbed heat, endothermic, ΔH positive. ✓



Extra practice 2

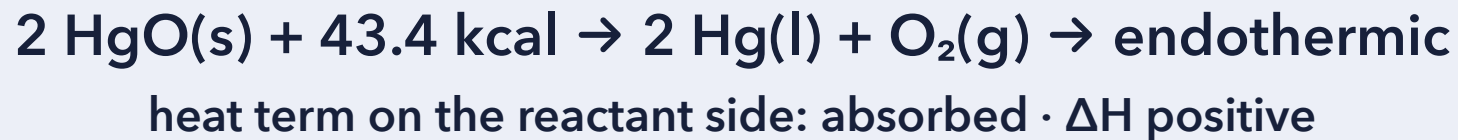


given: the heat term in kcal · 1 cal = 4.184 J

Strong heating breaks red mercury(II) oxide into mercury and oxygen. What is ΔH for the reaction, in kJ, sign included?

- A. +182
- B. -182
- C. +10.4
- D. +0.182
- E. +43.4

✓ Extra practice 2 · answer: A



Three conversion factors are needed: kcal $\rightarrow$ cal $\rightarrow$ J $\rightarrow$ kJ.

$$43.4 \text{ kcal} \times \frac{1000 \text{ cal}}{1 \text{ kcal}} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = +182 \text{ kJ} \cdot \text{answer A}$$

B read the heat term from the wrong side: it sits with the reactants, so it is absorbed, +182 kJ. C flipped the 4.184 factor: $43.4 \div 4.184 = 10.4$. D read kcal as cal: $43.4 \times 4.184 = 182$ J, only 0.182 kJ. E skipped the conversion: 43.4 is the heat in kcal, not kJ.

 **Sense check:** A kilojoule is smaller than a kilocalorie, so the kJ value must be the larger number: $43.4 \times 4.184 = 182$. ✓

Check yourself

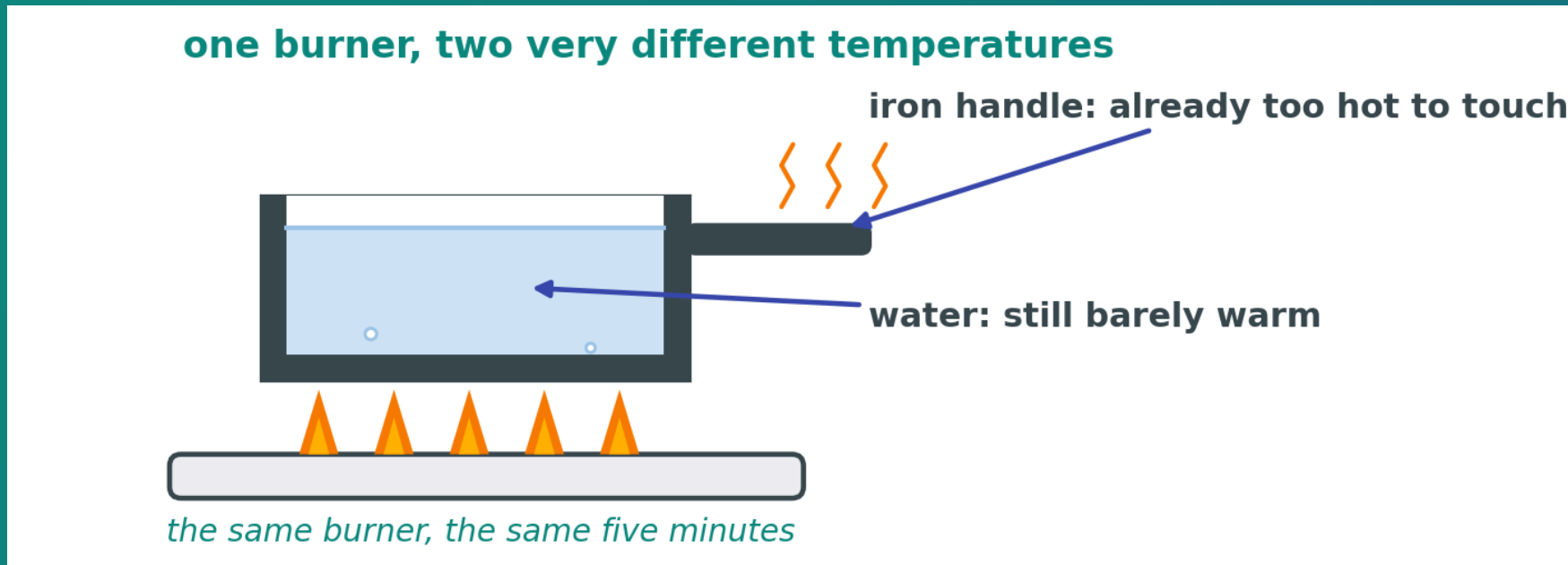
1. Propane burns in a camp stove: $\text{C}_3\text{H}_8(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 3 \text{CO}_2(\text{g}) + 4 \text{H}_2\text{O}(\text{l})$, $\Delta H = -2220$ kJ. State the verdict and which way heat flows.
2. An energy diagram shows the products 178 kJ above the reactants. Give the sign of ΔH and the verdict.

Exothermic or endothermic names the heat's direction. How much heat a sample gains or loses depends on its mass, its substance, and its temperature change: $q = m \cdot c \cdot \Delta T$.

3 · Specific Heat & $q = mc\Delta T$

Use $q = m \cdot c \cdot \Delta T$ to find the heat, the mass, the specific heat, or the temperature change, with ΔT measured final minus initial so the sign of q shows which way the heat flowed.

One burner, two temperatures



Five minutes in, the iron handle is too hot to touch; the water is barely warm. Iron needs far less heat than water for each degree it climbs.

Specific heat: joules per gram per degree

specific heat c : the heat that raises 1 g of a substance by 1 °C

J/g·°C: water 4.184 · ethyl alcohol 2.46 · aluminum 0.897 · iron 0.449 · copper 0.385 · gold 0.129 · lead 0.128

Heat flowing in raises a substance's temperature; heat flowing out lowers it. The joules needed to move each gram by one degree are fixed for each substance: its **specific heat**, c .

The heat equation

g and °C cancel between the factors, so q comes out in joules

$$\boxed{q} = \boxed{m} \times \boxed{c} \times \boxed{\Delta T}$$

heat in or out mass of the sample specific heat, the substance's property $T_{\text{final}} - T_{\text{initial}}$

J **g** **J/g·°C** **°C**

Three factors set the heat: the mass m , the substance's specific heat c , and the temperature change ΔT . One equation, four solvable quantities.

ΔT carries a sign

$$\Delta T = T_{\text{final}} - T_{\text{initial}}$$

heating 20.0 $\rightarrow$ 50.0 $^{\circ}\text{C}$: $\Delta T = +30.0$ $^{\circ}\text{C}$ · cooling 50.0 $\rightarrow$ 20.0 $^{\circ}\text{C}$: $\Delta T = -30.0$ $^{\circ}\text{C}$

ΔT is final minus initial, in that order. A cooling sample has a negative ΔT , so q comes out negative: the sample released heat. The sign records the direction of the flow.

Specific heat c vs heat capacity C

$$C = m \cdot c$$

60.0 g of water: $C = 60.0 \text{ g} \times 4.184 \text{ J/g}\cdot^\circ\text{C} = 251 \text{ J/}^\circ\text{C}$ · c stays $4.184 \text{ J/g}\cdot^\circ\text{C}$ for any amount of water

Specific heat is **intensive**: per gram, the same for a drop or a lake. Heat capacity C is **extensive**: it grows with the sample, because m is in it.

memory hook: little c , one gram · big C , the whole sample
the food Calorie's cue again: the capital letter marks the bigger quantity



The method

1. **List the pieces:** m , c , $\Delta T = T_{\text{final}} - T_{\text{initial}}$. Mark the unknown.
2. **Rearrange for the unknown** before numbers go in.
3. **Substitute and cancel units.**
4. **Check the sign:** cooling means negative ΔT and negative q .



Worked example 1: heat to warm water

$$q = m \cdot c \cdot \Delta T$$

given: 250 g water · $c = 4.184 \text{ J/g} \cdot ^\circ\text{C}$ · $20.0^\circ\text{C} \rightarrow 50.0^\circ\text{C}$ · wanted: q

A kettle warms **250 g of water** from 20.0°C to 50.0°C . How much heat does the water absorb? (c of water: $4.184 \text{ J/g} \cdot ^\circ\text{C}$)

List the pieces: m , c , and $\Delta T = T_{\text{final}} - T_{\text{initial}}$.

✓ Worked example 1: solution

$$q = m \cdot c \cdot \Delta T$$

given: 250 g water · c = 4.184 J/g·°C · 20.0 °C → 50.0 °C · wanted: q

Step 1 · List the pieces

Step 2 · Rearrange for the unknown

Step 3 · Substitute and cancel units

$$q = 250 \text{ g} \times \frac{4.184 \text{ J}}{1 \text{ g} \cdot \cancel{\text{°C}}} \times 30.0 \cancel{\text{°C}} = 31,400 \text{ J}$$

Step 4 · Check the sign

 **Sense check:** The water warmed, so ΔT and q are both positive: 31,400 J (31.4 kJ) absorbed. More grams or more degrees would cost more heat. ✓

✓ Worked example 1: solving for q

$$q = m \cdot c \cdot \Delta T$$

given: 250 g water · c = 4.184 J/g·°C · ΔT = +30.0 °C · found: q = 31,400 J

unknown

$$\boxed{q}_{\text{J}} = \boxed{m}_{\text{g}} \times \boxed{c}_{\text{J/g}\cdot^{\circ}\text{C}} \times \boxed{\Delta T}_{^{\circ}\text{C}}$$

q already stands alone: substitute and multiply the three factors

🧠 **Sense check:** q is the unknown, so no rearranging: m, c and ΔT go straight in. ✓

Worked example 2: find c , identify the substance

$$q = m \cdot c \cdot \Delta T$$

given: $q = 1347 \text{ J}$ · $m = 125 \text{ g}$ · $22.0 \text{ }^{\circ}\text{C} \rightarrow 46.0 \text{ }^{\circ}\text{C}$ · wanted: c

A **125-g** metal block absorbs **1347 J** as it warms from $22.0 \text{ }^{\circ}\text{C}$ to $46.0 \text{ }^{\circ}\text{C}$. Candidate specific heats, in $\text{J/g}\cdot^{\circ}\text{C}$:

aluminum	iron	copper	gold	lead
0.897	0.449	0.385	0.129	0.128

Find the block's specific heat and match it to the table.

✓ Worked example 2: solution

$$q = m \cdot c \cdot \Delta T$$

given: $q = 1347 \text{ J}$ · $m = 125 \text{ g}$ · $22.0 \text{ }^{\circ}\text{C} \rightarrow 46.0 \text{ }^{\circ}\text{C}$ · wanted: c

Step 1 · List the pieces

Step 2 · Rearrange for the unknown

$$q = m \cdot c \cdot \Delta T \rightarrow c = \frac{q}{m \cdot \Delta T}$$

Step 3 · Substitute and cancel units

$$c = \frac{1347 \text{ J}}{125 \text{ g} \times 24.0 \text{ }^{\circ}\text{C}} = 0.449 \text{ J/g}\cdot^{\circ}\text{C}$$

 **Sense check:** $0.449 \text{ J/g}\cdot^{\circ}\text{C}$ sits in the range of a metal's specific heat, ready to identify. ✓

✓ Worked example 2: identify the metal

$$c = 0.449 \text{ J/g}\cdot^{\circ}\text{C}$$

from $q = 1347 \text{ J} \cdot m = 125 \text{ g} \cdot \Delta T = +24.0^{\circ}\text{C}$, both positive

Match the property

aluminum	iron	copper	gold	lead
0.897	0.449	0.385	0.129	0.128

Step 4 • Check the sign

 **Sense check:** The block warmed: ΔT and q are both positive. Plug back in: $125 \text{ g} \times 0.449 \text{ J/g}\cdot^{\circ}\text{C} \times 24.0^{\circ}\text{C}$ returns 1347 J. ✓

✓ Worked example 2: solving for c

$$q = m \cdot c \cdot \Delta T$$

given: $q = 1347 \text{ J}$ · $m = 125 \text{ g}$ · $\Delta T = +24.0 \text{ }^{\circ}\text{C}$ · found: $c = 0.449 \text{ J/g}\cdot^{\circ}\text{C}$, iron

$$\boxed{q}_{\text{J}} = \boxed{m}_{\text{g}} \times \overset{\text{unknown}}{\boxed{c}_{\text{J/g}\cdot^{\circ}\text{C}}} \times \boxed{\Delta T}_{^{\circ}\text{C}}$$

↓ divide both sides by $m \cdot \Delta T$

$$\boxed{c} = \frac{\boxed{q}}{\boxed{m} \times \boxed{\Delta T}}$$

🧠 **Sense check:** c is the unknown. Dividing both sides by $m \cdot \Delta T$ isolates it before any numbers go in. ✓

Your turn: mass of ethyl alcohol

$$q = m \cdot c \cdot \Delta T$$

given: $q = 8020 \text{ J}$ · $c = 2.46 \text{ J/g} \cdot ^\circ\text{C}$ · $18.0^\circ\text{C} \rightarrow 43.0^\circ\text{C}$ · wanted: m

Ethyl alcohol ($c = 2.46 \text{ J/g} \cdot ^\circ\text{C}$) absorbs **8020 J** and warms from 18.0°C to 43.0°C .

$$m = \frac{q}{c \cdot \Delta T} = \frac{8020 \text{ J}}{\text{ } \text{J/g} \cdot ^\circ\text{C} \times \text{ } ^\circ\text{C}} = \text{ } \text{g}$$

Fill in c and $\Delta T = T_{\text{final}} - T_{\text{initial}}$, then compute the mass.

$$m = \frac{8020 \text{ J}}{2.46 \text{ J/g} \cdot ^\circ\text{C} \times 25.0 ^\circ\text{C}} = 130. \text{ g}$$

 **Sense check:** 2.46 J warms one gram by one degree, so 8020 J spread over 25.0 degrees warms about 130 g . ✓

⚠ Where this goes wrong

$$q = m \cdot c \cdot \Delta T$$

$$250 \text{ g water} \cdot c = 4.184 \text{ J/g}\cdot^{\circ}\text{C} \cdot 20.0 \rightarrow 50.0^{\circ}\text{C} \cdot \text{correct } q = 31,400 \text{ J}$$

Leaving out the mass. $4.184 \times 30.0 = 126 \text{ J}$ is the heat for a single gram. The sample has 250 of them. All three factors multiply: $q = m \cdot c \cdot \Delta T = 31,400 \text{ J}$.

Leaving out the temperature change. $250 \times 4.184 = 1046 \text{ J}$ warms the water by one degree only. Multiply by the full ΔT of 30.0°C .

Dividing by the specific heat. $250 \times 30.0 \div 4.184 = 1790$, and its units are $\text{g}^2\cdot^{\circ}\text{C}^2/\text{J}$, not joules. c multiplies on top: $(4.184 \text{ J} / 1 \text{ g}\cdot^{\circ}\text{C})$.

Subtracting the temperatures in the wrong order. $\Delta T = 20.0 - 50.0 = -30.0^{\circ}\text{C}$ gives $q = -31,400 \text{ J}$: heat released by water that is warming. ΔT is $T_{\text{final}} - T_{\text{initial}}$.

 Practice 1

$$q = m \cdot c \cdot \Delta T$$

given: 75.0 g copper · $c = 0.385 \text{ J/g}\cdot^\circ\text{C}$ · $\Delta T = +20.0^\circ\text{C}$ · wanted: q

A **75.0-g copper** fitting warms by **20.0°C** as hot water flows past it. How much heat, in J, does the copper absorb? (*c of copper: $0.385 \text{ J/g}\cdot^\circ\text{C}$*)

A. 7.70

B. 28.9

C. 578

D. 3.90×10^3

✓ Practice 1 · answer: C

$$q = m \cdot c \cdot \Delta T$$

$$\text{given: } 75.0 \text{ g copper} \cdot c = 0.385 \text{ J/g} \cdot ^\circ\text{C} \cdot \Delta T = +20.0 ^\circ\text{C}$$

$$q = 75.0 \text{ g} \times \frac{0.385 \text{ J}}{1 \text{ g} \cdot ^\circ\text{C}} \times 20.0 ^\circ\text{C} = \mathbf{578 \text{ J (answer C)}}$$

A left out the mass: $0.385 \times 20.0 = 7.70 \text{ J}$ warms one gram. B left out the temperature change: $75.0 \times 0.385 = 28.9 \text{ J}$ is one degree's worth. D divided by the specific heat: $75.0 \times 20.0 \div 0.385 = 3.90 \times 10^3$, with units $\text{g}^2 \cdot ^\circ\text{C}^2/\text{J}$.

 **Sense check:** Copper takes only 0.385 J per gram per degree, but 75 grams and 20 degrees multiply that into hundreds of joules. ✓

Worked example 3: final temperature of a cooling sample

$$q = m \cdot c \cdot \Delta T$$

given: 150 g aluminum · $c = 0.897 \text{ J/g} \cdot ^\circ\text{C}$ · $T_{\text{initial}} = 95.0 \text{ }^\circ\text{C}$ · releases 8100 J · wanted: T_{final}

A **150-g aluminum** pan lid at **$95.0 \text{ }^\circ\text{C}$** releases **8100 J** as it cools. What is its final temperature? (*c of aluminum: $0.897 \text{ J/g} \cdot ^\circ\text{C}$*)

A common first attempt: substitute 8100 J with no sign. Test the result.

✓ Worked example 3: solution

$$q = m \cdot c \cdot \Delta T$$

given: 150 g aluminum · c = 0.897 J/g·°C · T_{initial} = 95.0 °C · releases 8100 J · wanted: T_{final}

A common first attempt

$$\Delta T = \frac{+8100 \text{ J}}{150 \text{ g} \times 0.897 \text{ J/g} \cdot ^\circ\text{C}} = +60.2 \text{ } ^\circ\text{C} \rightarrow T_{\text{final}} = 95.0 + 60.2 = \mathbf{155.2 \text{ } ^\circ\text{C} \times}$$

Step 1 · List the pieces

Step 2 · Rearrange for the unknown

$$\Delta T = \frac{q}{m \cdot c}, \text{ then } T_{\text{final}} = T_{\text{initial}} + \Delta T$$

🧠 **Sense check:** With q entered as −8100 J, the formula is set to return a negative ΔT: a temperature drop. ✓

✓ Worked example 3: final temperature

$$q = -8100 \text{ J released} \cdot 150 \text{ g aluminum} \cdot c = 0.897 \text{ J/g}\cdot^{\circ}\text{C}$$

$$T_{\text{initial}} = 95.0^{\circ}\text{C} \cdot \text{wanted: } T_{\text{final}}$$

Step 3 • Substitute and cancel units

$$\Delta T = \frac{-8100 \text{ J}}{150 \text{ g} \times 0.897 \text{ J/g}\cdot^{\circ}\text{C}} = \frac{-8100 \text{ J}}{134.55 \text{ J/}^{\circ}\text{C}} = -60.2^{\circ}\text{C}$$

$$T_{\text{final}} = 95.0^{\circ}\text{C} + (-60.2^{\circ}\text{C}) = \mathbf{34.8^{\circ}\text{C}}$$

Step 4 • Check the sign

 **Sense check:** Released heat means negative q , negative ΔT , and a lower final temperature: $95.0 \rightarrow 34.8^{\circ}\text{C}$. ✓ The signless route predicted 155.2°C , a cooling lid ending hotter. ✗

✓ Worked example 3: solving for ΔT , then T_{final}

$$q = -8100 \text{ J} \cdot 150 \text{ g aluminum} \cdot c = 0.897 \text{ J/g} \cdot ^\circ\text{C} \cdot T_{\text{initial}} = 95.0 \text{ } ^\circ\text{C}$$

$$\text{found: } \Delta T = -60.2 \text{ } ^\circ\text{C} \cdot T_{\text{final}} = 34.8 \text{ } ^\circ\text{C}$$

$$\begin{array}{c} \boxed{q} = \boxed{m} \times \boxed{c} \times \boxed{\Delta T} \\ \text{J} \quad \text{g} \quad \text{J/g} \cdot ^\circ\text{C} \quad ^\circ\text{C} \end{array}$$

↓ divide both sides by $m \cdot c$

$$\boxed{\Delta T} = \frac{\boxed{q}}{\boxed{m} \times \boxed{c}} \quad \text{then} \quad \boxed{T_{\text{final}} = T_{\text{initial}} + \Delta T}$$

🧠 **Sense check:** ΔT is a change, not an end point. $T_{\text{final}} = 95.0 + (-60.2) = 34.8 \text{ } ^\circ\text{C}$. ✓

Take-home: ΔT is final minus initial

warming: $20.0\text{ }^{\circ}\text{C} \rightarrow 50.0\text{ }^{\circ}\text{C} \cdot \Delta T = 50.0 - 20.0 = +30.0\text{ }^{\circ}\text{C} \cdot q$ positive
heat absorbed ✓

cooling: $95.0\text{ }^{\circ}\text{C} \rightarrow 34.8\text{ }^{\circ}\text{C} \cdot \Delta T = 34.8 - 95.0 = -60.2\text{ }^{\circ}\text{C} \cdot q$ negative
heat released ✓

ΔT is always $T_{\text{final}} - T_{\text{initial}}$, and released heat enters as negative q . The sign is part of the quantity; it records which way the heat flowed.

 Practice 2

$$q = m \cdot c \cdot \Delta T$$

given: 100. g lead $\cdot$ $c = 0.128 \text{ J/g}\cdot^\circ\text{C}$ $\cdot$ $62.0^\circ\text{C} \rightarrow 22.0^\circ\text{C}$ $\cdot$ wanted: q

A **100.-g lead** sinker at 62.0°C drops into a stream and cools to 22.0°C . What is q for the lead, in J? (*c of lead: $0.128 \text{ J/g}\cdot^\circ\text{C}$*)

- A. 512
- B. -512
- C. 12.8
- D. -5.12

✓ Practice 2 · answer: B

$$q = m \cdot c \cdot \Delta T$$

$$\text{given: } 100. \text{ g lead} \cdot c = 0.128 \text{ J/g} \cdot ^\circ\text{C} \cdot \Delta T = 22.0 - 62.0 = -40.0 \text{ } ^\circ\text{C}$$

$$q = 100. \text{ g} \times \frac{0.128 \text{ J}}{1 \text{ g} \cdot ^\circ\text{C}} \times (-40.0 \text{ } ^\circ\text{C}) = -512 \text{ J (answer B)}$$

A subtracted the temperatures in the wrong order: $62.0 - 22.0 = +40.0 \text{ } ^\circ\text{C}$ gives $+512 \text{ J}$, heat absorbed by a cooling sinker. C stopped at $m \times c$: $100. \times 0.128 = 12.8 \text{ J}$, one degree's worth with no sign. D left out the mass: $0.128 \times (-40.0) = -5.12 \text{ J}$, the heat for a single gram.

 **Sense check:** The sinker cooled 40 degrees, so it released heat: q must be negative. ✓



Practice 3

$$q = m \cdot c \cdot \Delta T$$

given: 45.0 g water · $q = +2260 \text{ J}$ · $c = 4.184 \text{ J/g}\cdot^{\circ}\text{C}$ · wanted: ΔT

45.0 g of water absorbs **2260 J** of heat. By how much does its temperature rise, in $^{\circ}\text{C}$? (c of water: $4.184 \text{ J/g}\cdot^{\circ}\text{C}$)

- A. 0.0833
- B. 50.2
- C. 12.0
- D. 210.
- E. 540.

✓ Practice 3 · answer: C

$$\Delta T = q / (m \cdot c)$$

$$\text{given: } 45.0 \text{ g water} \cdot q = +2260 \text{ J} \cdot c = 4.184 \text{ J/g}\cdot^{\circ}\text{C}$$

$$\Delta T = \frac{2260 \text{ J}}{45.0 \text{ g} \times 4.184 \text{ J/g}\cdot^{\circ}\text{C}} = +12.0^{\circ}\text{C} \text{ (answer C)}$$

A flipped the rearrangement: $45.0 \times 4.184 \div 2260 = 0.0833^{\circ}\text{C}$. B left out c: $2260 \div 45.0 = 50.2^{\circ}\text{C}$. D multiplied by c: $2260 \times 4.184 \div 45.0 = 210.^{\circ}\text{C}$. E left out the mass: $2260 \div 4.184 = 540.^{\circ}\text{C}$, the change for a single gram.

🧠 **Sense check:** Substitute back: $45.0 \text{ g} \times 4.184 \text{ J/g}\cdot^{\circ}\text{C} \times 12.0^{\circ}\text{C} = 2260 \text{ J}$. ✓



Practice 4

$$q = m \cdot c \cdot \Delta T$$

given: 25.0 g gold · $T_{\text{initial}} = 27.0\text{ }^{\circ}\text{C}$ · absorbs 2.34 kJ · $c = 0.129\text{ J/g}\cdot^{\circ}\text{C}$ · wanted: T_{final}

25.0 g of gold at $27.0\text{ }^{\circ}\text{C}$ absorbs **2.34 kJ** of heat. What is the final temperature, in $^{\circ}\text{C}$? (*c of gold: $0.129\text{ J/g}\cdot^{\circ}\text{C}$*)

- A. 27.7
- B. 699
- C. 726
- D. 753
- E. 39.1

✓ Practice 4 · answer: D

$$\Delta T = q / (m \cdot c), \text{ then } T_{\text{final}} = T_{\text{initial}} + \Delta T$$

$$\text{given: } 25.0 \text{ g gold} \cdot T_{\text{initial}} = 27.0 \text{ }^{\circ}\text{C} \cdot q = +2340 \text{ J} \cdot c = 0.129 \text{ J/g}\cdot^{\circ}\text{C}$$

$$\Delta T = \frac{2340 \text{ J}}{25.0 \text{ g} \times 0.129 \text{ J/g}\cdot^{\circ}\text{C}} = +725.6 \text{ }^{\circ}\text{C} \rightarrow T_{\text{final}} = 27.0 + 725.6 = \mathbf{753 \text{ }^{\circ}\text{C (answer D)}}$$

A left q in kilojoules: $\Delta T = 0.726 \text{ }^{\circ}\text{C}$, so $27.7 \text{ }^{\circ}\text{C}$. B subtracted T_{initial} : $725.6 - 27.0 = 699 \text{ }^{\circ}\text{C}$; $T_{\text{final}} = T_{\text{initial}} + \Delta T$. C stopped at ΔT : $726 \text{ }^{\circ}\text{C}$ is the change, not the final temperature. E multiplied by c : $2340 \times 0.129 \div 25.0 = 12.07 \text{ }^{\circ}\text{C}$, then $39.1 \text{ }^{\circ}\text{C}$.

 **Sense check:** A tiny specific heat means a huge swing: over 700 degrees, most of the way to gold's $1064 \text{ }^{\circ}\text{C}$ melting point. ✓



Practice 5

$$q = m \cdot c \cdot \Delta T$$

given: 2.0 kg aluminum pan · 23.0 °C → 180.0 °C · $c = 0.89 \text{ J/g} \cdot ^\circ\text{C}$ · wanted: q · the mass arrives in kilograms

A **2.0-kg aluminum** pan heats from 23.0 °C to 180.0 °C on a burner. How much heat, in J, does the pan absorb? (*c of aluminum: $0.89 \text{ J/g} \cdot ^\circ\text{C}$*)

- A. 280
- B. 1.8×10^3
- C. 2.8×10^5
- D. 3.2×10^5

✓ Practice 5 · answer: C

$$q = m \cdot c \cdot \Delta T$$

$$m = 2.0 \text{ kg} = 2000 \text{ g} \cdot c = 0.89 \text{ J/g} \cdot ^\circ\text{C} \cdot \Delta T = 180.0 - 23.0 = +157.0 ^\circ\text{C}$$

$$q = 2.0 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{0.89 \text{ J}}{1 \text{ g} \cdot ^\circ\text{C}} \times 157.0 ^\circ\text{C} = 279,460 \text{ J} \approx \mathbf{2.8 \times 10^5 \text{ J (answer C)}}$$

c is per gram, so the kilograms convert inside the chain: 2.0 kg is 2000 g. A used 2.0 as if it were grams: $2.0 \times 0.89 \times 157.0 = 280 \text{ J}$, a thousand times short. B left out the temperature change: $2000 \times 0.89 = 1.8 \times 10^3 \text{ J}$, one degree's worth. D used 180.0 as ΔT : $2000 \times 0.89 \times 180.0 = 3.2 \times 10^5 \text{ J}$. ΔT is final minus initial: $157.0 ^\circ\text{C}$.

 **Sense check:** Two thousand grams climbing 157 degrees, even at aluminum's modest 0.89 J per gram per degree, costs about 280 kJ. ✓



Check yourself

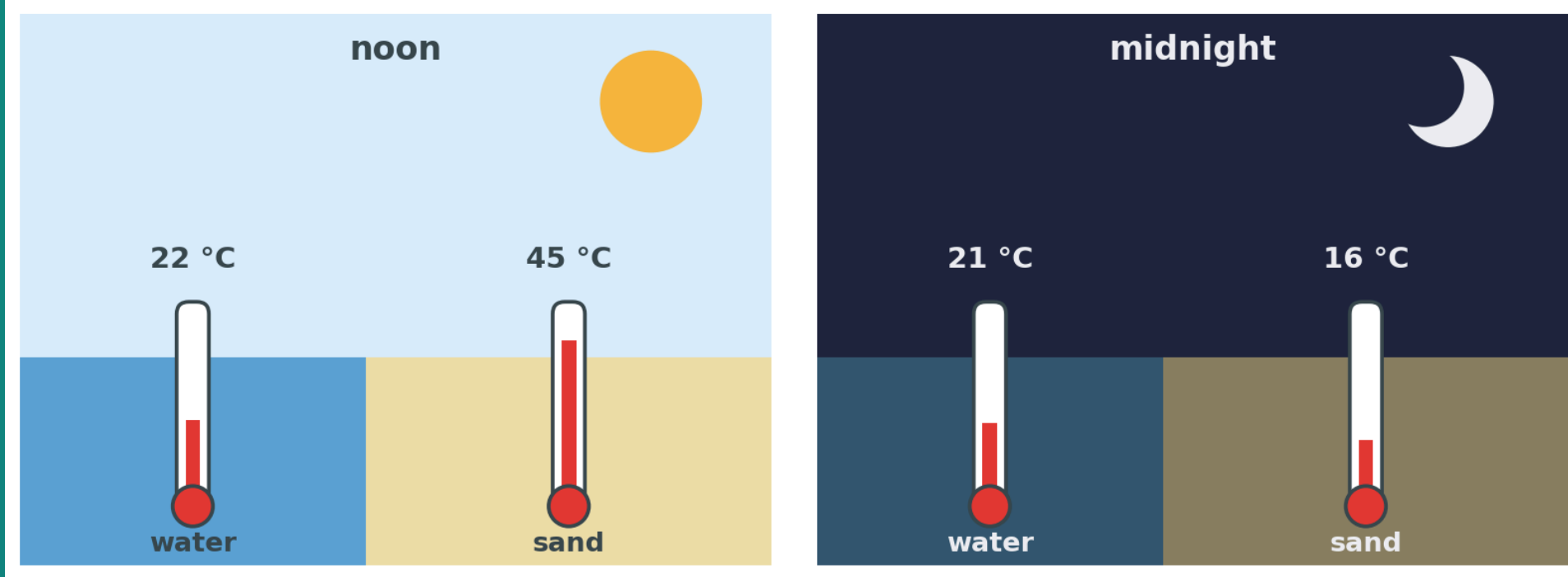
1. Water cools from 50.0 °C to 20.0 °C. Write ΔT with its sign. What is the sign of q , and what does it say about the heat?
2. Solve $q = m \cdot c \cdot \Delta T$ for m , in symbols. Which units cancel, and which unit survives?

Two samples can receive the same heat and change temperature by different amounts. Equal masses, equal q : the substance with the smaller specific heat shows the larger ΔT .

4 · Comparing Specific Heats

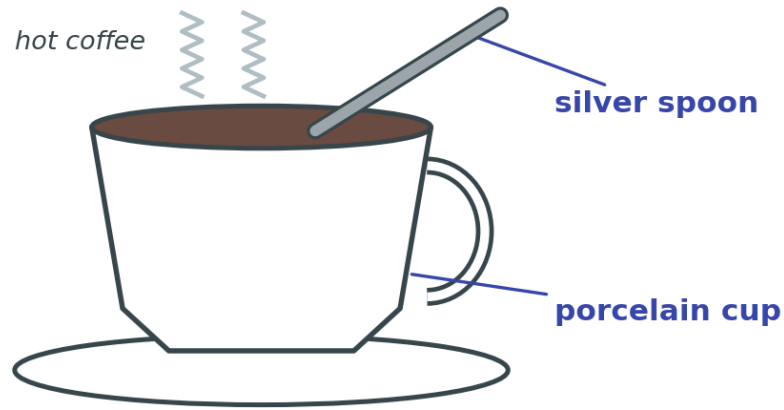
For the same heat, read specific heat in reverse (equal masses, the smaller c means the larger temperature change; unequal masses, compare the whole sample's $m \cdot c$), confirming with $\Delta T = q/mc$ when a number is wanted.

The beach, twelve hours apart



At noon the sand scorches bare feet while the ocean stays cool. By midnight the sand is cold, and the water is the warm place to be.

Silver spoon, porcelain cup



Specific heat, c : the heat that warms 1 g by 1 °C. Which has the higher c , the spoon or the cup?

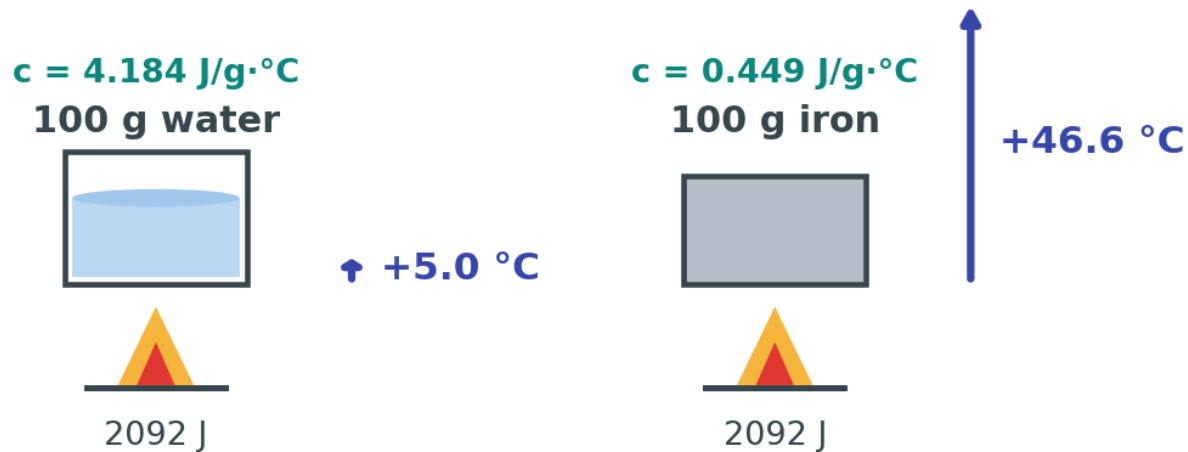
The cup. Per gram, porcelain needs about four times the heat per degree.

silver: $c = 0.235 \text{ J/g}\cdot^\circ\text{C}$ · porcelain: $c \approx 1 \text{ J/g}\cdot^\circ\text{C}$

the same heat per gram swings the spoon's temperature about four times as far

Same heat and mass: c and ΔT trade off

the same heat into the same mass

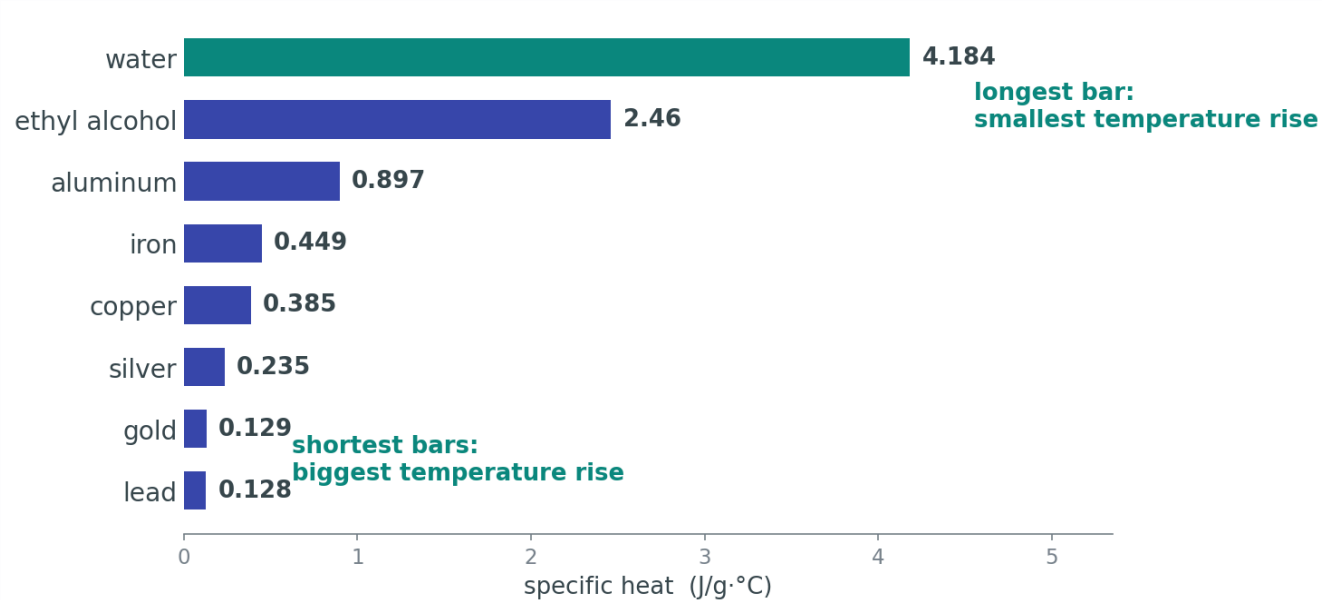


Give equal masses the same heat. Their temperature changes are not equal. Specific heat sits in the denominator of $\Delta T = q/mc$, so the smaller c , the larger the temperature change.

$$\Delta T = q / (m \cdot c)$$

same q , same m : c in the denominator: smaller c , larger ΔT

Ranking substances by specific heat



Every substance has its own specific heat; water's is several times any metal's. Higher c , smaller temperature change from the same heat.

water 4.184 · ethyl alcohol 2.46 · aluminum 0.897 · iron 0.449 · copper 0.385 · silver 0.235 · gold 0.129 · lead 0.128

specific heat c , in J/g·°C: smaller c , larger temperature change for the same heat and mass

Why water resists temperature swings

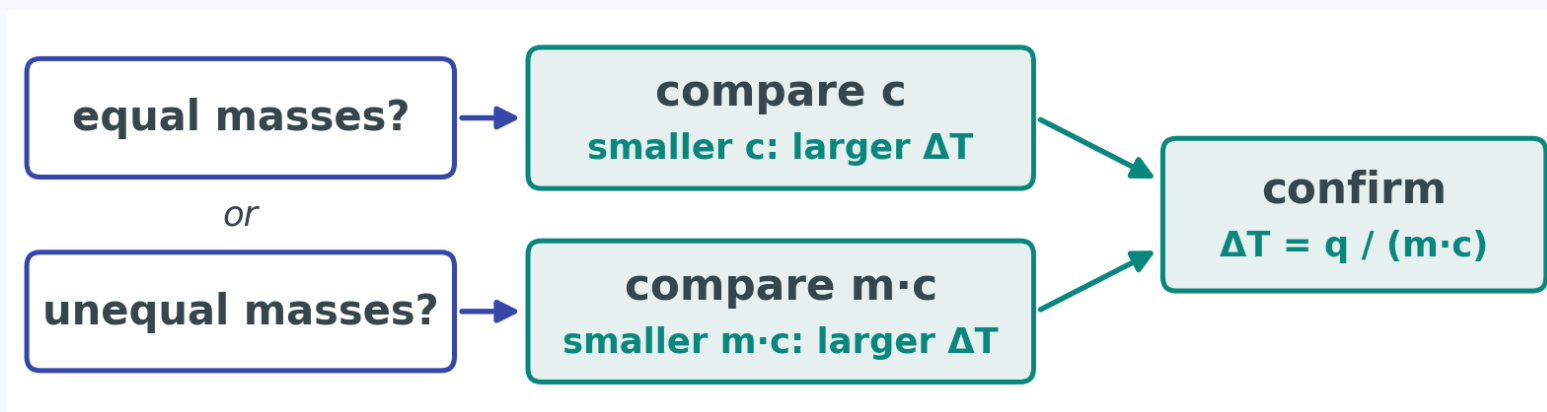
Water soaks up heat with only a small temperature rise and releases it slowly, so coastlines stay mild. The cause is the strong grip between water molecules: hydrogen bonding.

2092 J into 100 g of each: water rises +5.0 °C · iron rises +46.6 °C
water's specific heat is 9.3× iron's, so the same heat moves it 9.3× less



The method

1. **Compare the specific heats.** Same heat and mass: smaller c means larger ΔT .
2. **Name the response.** The substance with the smaller c swings more.
3. **Confirm with $\Delta T = q/mc$.** c and m are in the denominator.





Guided example: water and gold

$$\Delta T = q / (m \cdot c)$$

10.0 g water ($c = 4.184$) and 10.0 g gold ($c = 0.129$) $\cdot q = 100. \text{ J each} \cdot$ wanted: each ΔT

10.0 g of water and **10.0 g of gold** each absorb **100. J**. Which changes temperature more, and by how much?

The masses are equal, so start with the specific heats.

✓ Guided example: solution

$$\Delta T = q / (m \cdot c)$$

$$10.0 \text{ g water } (c = 4.184) \cdot 10.0 \text{ g gold } (c = 0.129) \cdot q = 100. \text{ J each}$$

Step 1 · Compare the specific heats

Step 2 · Name the response

Step 3 · Confirm with $\Delta T = q/mc$

$$\text{water: } \Delta T = \frac{100. \text{ J}}{10.0 \text{ g} \times 4.184 \text{ J/g} \cdot ^\circ\text{C}} = +2.39 \text{ } ^\circ\text{C}$$

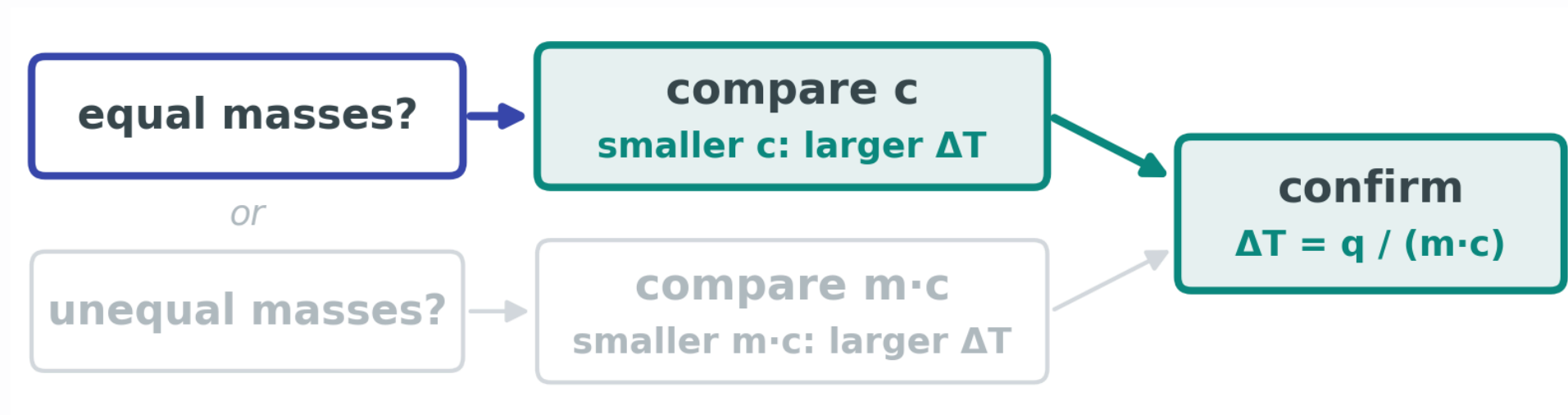
$$\text{gold: } \Delta T = \frac{100. \text{ J}}{10.0 \text{ g} \times 0.129 \text{ J/g} \cdot ^\circ\text{C}} = \mathbf{+77.5 \text{ } ^\circ\text{C}}$$

 **Sense check:** Gold rises 77.5 °C, water only 2.39 °C. Water's c is 32 times gold's, so gold swings 32 times as far. ✓

✓ Guided example: the route on the strip

$$\Delta T = q / (m \cdot c)$$

given: 10.0 g water · 10.0 g gold · 100. J each · found: water +2.39 °C · gold +77.5 °C



🧠 **Sense check:** Equal masses, so c alone decides which swings more. The equation confirms it. ✓

Practice 1 · which ends hotter

$$\Delta T = q / (m \cdot c)$$

aluminum: $c = 0.897 \text{ J/g}\cdot^{\circ}\text{C}$ · silver: $c = 0.235 \text{ J/g}\cdot^{\circ}\text{C}$

Equal masses of aluminum and silver start at 25°C and absorb the same heat. Which ends hotter?

- A. Aluminum
- B. Silver
- C. Both end at the same temperature
- D. It depends on how much heat they absorb

✓ Practice 1 · answer: B

$$\Delta T = q / (m \cdot c) \cdot \text{equal masses, same heat: compare } c$$

silver 0.235 · aluminum 0.897 J/g·°C · smaller c, larger ΔT

A read specific heat backwards: the larger c gives the smaller rise. C equal heat gives equal ΔT only when the specific heats match. D the heat sets how far each one warms, not which warms more: the smaller c wins at any heat.

🧠 **Sense check:** Silver ends hotter. Its temperature swings about 3.8 times as far as aluminum's (0.897 / 0.235). ✓

Practice 2 · one temperature change

$$\Delta T = q / (m \cdot c)$$

given: 30.0 g iron ($c = 0.449$) · 30.0 g aluminum ($c = 0.897$) · 450. J each

A **30.0 g** iron bolt and a **30.0 g** aluminum bolt each absorb **450. J**. What is the iron bolt's temperature change, in °C?

- A. 16.7
- B. 6.74
- C. 15.0
- D. 33.4

✓ Practice 2 · answer: D

$$\Delta T = q / (m \cdot c)$$

$$\text{iron: } m = 30.0 \text{ g} \cdot c = 0.449 \text{ J/g} \cdot ^\circ\text{C} \cdot q = 450. \text{ J}$$

$$\text{iron: } \Delta T = \frac{450. \text{ J}}{30.0 \text{ g} \times 0.449 \text{ J/g} \cdot ^\circ\text{C}} = \frac{450. \cancel{\text{J}}}{13.47 \cancel{\text{J}}/^\circ\text{C}} = +33.4 \text{ } ^\circ\text{C} \text{ (answer D)}$$

A used aluminum's c: $450. / (30.0 \times 0.897) = 16.7$. B multiplied by c instead of dividing: $450. \times 0.449 / 30.0 = 6.74$. C dropped c: $450. / 30.0 = 15.0$.

🧠 **Sense check:** Iron's c is half of aluminum's, so iron swings twice as far: $33.4 \text{ } ^\circ\text{C}$ against $16.7 \text{ } ^\circ\text{C}$. ✓



Practice 3 · ranking three substances

$$\Delta T = q / (m \cdot c)$$

given: iron $c = 0.449$ · ethyl alcohol $c = 2.46$ · copper $c = 0.385 \text{ J/g}\cdot^\circ\text{C}$ · equal masses · same heat

Equal masses of iron, ethyl alcohol and copper absorb the same heat. Rank them by ΔT , largest first.

- A. ethyl alcohol > iron > copper
- B. copper = iron > ethyl alcohol
- C. copper > iron > ethyl alcohol
- D. All three change by the same amount

✓ Practice 3 · answer: C

$\Delta T = q / (m \cdot c)$ · equal masses, same heat: the smallest c swings most
copper 0.385 < iron 0.449 < ethyl alcohol 2.46 J/g·°C

A ranked by c itself: the largest c gives the smallest ΔT , not the largest. B lumped the metals together: copper's 0.385 sits below iron's 0.449, so copper still swings further. D equal heat gives equal ΔT only when the specific heats match.

🧠 **Sense check:** Copper swings a little further than iron, and about 6.4 times as far as ethyl alcohol (2.46 / 0.385). ✓



Worked example: when the masses differ

$$\Delta T = q / (m \cdot c)$$

200 g copper ($c = 0.385$) and 50.0 g aluminum ($c = 0.897$) · $q = 2000$ J each · wanted: which ends hotter

A **200 g copper** block and a **50.0 g aluminum** block each absorb **2000 J**.

A common first answer: copper has the smaller specific heat, so copper wins. Check it: the masses are not equal.

✓ Worked example: solution

$$\Delta T = q / (m \cdot c)$$

200 g copper ($c = 0.385$) · 50.0 g aluminum ($c = 0.897$) · $q = 2000$ J each

Step 1 · Compare the specific heats

Step 2 · Name the response

Step 3 · Confirm with $\Delta T = q/mc$

$$\text{copper: } \Delta T = \frac{2000 \text{ J}}{200 \text{ g} \times 0.385 \text{ J/g} \cdot ^\circ\text{C}} = \frac{2000 \text{ J}}{77.0 \text{ J/}^\circ\text{C}} = +26.0 \text{ }^\circ\text{C}$$

$$\text{aluminum: } \Delta T = \frac{2000 \text{ J}}{50.0 \text{ g} \times 0.897 \text{ J/g} \cdot ^\circ\text{C}} = \frac{2000 \text{ J}}{44.85 \text{ J/}^\circ\text{C}} = +44.6 \text{ }^\circ\text{C}$$

 **Sense check:** Aluminum ends hotter, 44.6 °C against 26.0 °C, even though copper has the smaller specific heat. When the masses differ, compare $m \cdot c$, not c alone. ✓

✓ Worked example: the route on the strip

$$\Delta T = q / (m \cdot c)$$

given: 200 g copper · 50.0 g aluminum · 2000 J each · found: copper +26.0 °C · aluminum +44.6 °C



 **Sense check:** Unequal masses, so compare $m \cdot c$: aluminum's 44.85 J/°C is the smaller, so aluminum swings more. The equation confirms it. ✓

Take-home: compare $m \cdot c$, not c alone

200 g copper: $m \cdot c = 77.0 \text{ J/}^\circ\text{C}$ · 50.0 g aluminum: $m \cdot c = 44.85 \text{ J/}^\circ\text{C}$
same 2000 J: copper $+26.0 \text{ }^\circ\text{C}$ · aluminum $+44.6 \text{ }^\circ\text{C}$ · the smaller $m \cdot c$ swings more

Comparing c alone assumes equal masses. When the masses differ, the whole sample's $m \cdot c$ sets the response: the smaller $m \cdot c$, the larger the temperature change.

Where this goes wrong

Equal heat, equal temperature. Equal q into equal mass does not give equal ΔT . 1000 J into 50.0 g of water raises it 4.78 °C; the same 1000 J into 50.0 g of copper raises it 51.9 °C. The specific heats differ, so the temperature changes differ.

Right substance, backwards reason. Naming the low- c substance as the one that ends hotter "because it stores more heat per gram." It stores less per gram: lead takes 0.128 J to warm a gram by a degree, aluminum 0.897 J. That low cost per degree is exactly why lead swings more.

Comparing c when the masses differ. The rule "smaller c wins" assumes equal mass. When the masses are not equal, mass is in the denominator too. Compare the whole sample's $m \cdot c$, or run $\Delta T = q/mc$ for each.



Practice 4

$$\Delta T = q / (m \cdot c)$$

given: 400. g lead ($c = 0.128$) · 150. g silver ($c = 0.235$) · 1200 J each · same start

A **400. g block of lead** ($c = 0.128 \text{ J/g}\cdot^\circ\text{C}$) and a **150. g block of silver** ($c = 0.235 \text{ J/g}\cdot^\circ\text{C}$) start at the same temperature and each absorb **1200 J**. Which statement matches the outcome?

- A. Lead warms by about 23.4°C and silver by about 34.0°C , so silver ends hotter.
- B. Lead ends hotter: its smaller specific heat turns the same heat into the larger rise.
- C. Both warm by the same amount, since each absorbs the same 1200 J.
- D. Lead warms by about 34.0°C and silver by about 23.4°C , so lead ends hotter.

✓ Practice 4 · answer: A

$\Delta T = q / (m \cdot c)$ · the masses differ, so compare $m \cdot c$

lead: $400. \times 0.128 = 51.2 \text{ J/}^\circ\text{C}$ · silver: $150. \times 0.235 = 35.25 \text{ J/}^\circ\text{C}$ · silver's is smaller

$$\text{lead: } \Delta T = \frac{1200 \text{ J}}{51.2 \text{ J/}^\circ\text{C}} = +23.4 \text{ }^\circ\text{C} \cdot \text{silver: } \Delta T = \frac{1200 \text{ J}}{35.25 \text{ J/}^\circ\text{C}} = +34.0 \text{ }^\circ\text{C (answer A)}$$

B compared c alone: "smaller c wins" holds only at equal masses, and the lead block is heavier. C equal heat does not mean equal ΔT : the two samples' $m \cdot c$ differ. D computed both rises and pinned them on the wrong metals.

🧠 **Sense check:** Silver has the larger c, yet its smaller $m \cdot c$ makes it swing more: $34.0 \text{ }^\circ\text{C}$ against $23.4 \text{ }^\circ\text{C}$. ✓



Extra practice 1

$$\Delta T = q / (m \cdot c)$$

lead: $c = 0.128 \text{ J/g}\cdot^\circ\text{C}$ · copper: $c = 0.385 \text{ J/g}\cdot^\circ\text{C}$ · equal masses · same heat

Two blocks of equal mass, one lead and one copper, take in the same heat. The lead warms by **36.0 °C**. By how much, in °C, does the copper warm?

- A. 36.0
- B. 12.0
- C. 108
- D. 4.61

✓ Extra practice 1 · answer: B

$\Delta T = q / (m \cdot c)$ · equal masses, same heat: c alone decides
lead 0.128 · copper 0.385 J/g·°C · copper's c is larger, so copper warms less

$$q/m = 0.128 \times 36.0 = 4.61 \text{ J/g} \rightarrow \text{copper: } \Delta T = \frac{4.61 \text{ J/g}}{0.385 \text{ J/g} \cdot ^\circ\text{C}} = +12.0 ^\circ\text{C} \cdot \text{answer B}$$

A took equal heat as an equal rise: 36.0 °C holds only when the specific heats match. C flipped the ratio: $36.0 \times 0.385 / 0.128 = 108$, but the larger c gives the smaller rise. D stopped halfway: $0.128 \times 36.0 = 4.61$ is the heat per gram, in J/g, not copper's rise.

🧠 **Sense check:** Copper's c is 3.0 times lead's ($0.385 / 0.128$), so copper warms a third as far: $36.0 / 3.0 = 12.0 ^\circ\text{C}$. ✓



Extra practice 2

$$q = m \cdot c \cdot \Delta T$$

$$\text{silver: } c = 0.235 \text{ J/g}\cdot^{\circ}\text{C} \cdot \text{iron: } c = 0.449 \text{ J/g}\cdot^{\circ}\text{C} \cdot 1 \text{ cal} = 4.184 \text{ J}$$

A silver bar and an iron bar, **60.0 g** each, both warm by **25.0 °C**. How much more heat, in cal, does the iron bar absorb?

- A. 76.7
- B. 321
- C. 1.34×10^3
- D. 161

✓ Extra practice 2 · answer: A

$q = m \cdot c \cdot \Delta T$ · same m and ΔT : the larger c takes more heat
 60.0 g each · +25.0 °C each · iron 0.449 · silver 0.235 J/g·°C · 1 cal = 4.184 J

iron: $60.0 \times 0.449 \times 25.0 = 673.5 \text{ J}$ · silver: $60.0 \times 0.235 \times 25.0 = 352.5 \text{ J}$

$$673.5 \text{ J} - 352.5 \text{ J} = 321 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = \mathbf{76.7 \text{ cal} \cdot \text{answer A}}$$

B stopped in joules: 321 J is the right gap in the wrong unit. C put the factor upside down: $321 \times 4.184 = 1.34 \times 10^3$. D never subtracted silver's heat: $673.5 / 4.184 = 161 \text{ cal}$ is iron's whole heat.

 **Sense check:** Iron's c beats silver's by 0.214 J/g·°C. Over 60.0 g and 25.0 °C that is 321 J, about 77 cal. ✓



Extra practice 3

$$\Delta T = q / (m \cdot c), \text{ then } T_{\text{final}} = T_{\text{initial}} + \Delta T$$

iron: $c = 0.449 \text{ J/g}\cdot^{\circ}\text{C}$ · aluminum: $c = 0.897 \text{ J/g}\cdot^{\circ}\text{C}$

A **140. g iron block** at **30.0 °C** and a **40.0 g aluminum block** at **22.0 °C** each absorb **1500. J**. Which statement matches the outcome?

- A. Iron ends hotter: it starts warmer, and the same heat gives the same rise.
- B. Iron ends hotter, at 53.9 °C: its smaller specific heat gives the larger rise.
- C. Aluminum ends hotter, at 41.8 °C.
- D. Aluminum ends hotter, at 63.8 °C.

✓ Extra practice 3 · answer: D

$\Delta T = q / (m \cdot c)$ · the masses differ, so compare $m \cdot c$

iron: $140. \times 0.449 = 62.86 \text{ J/}^\circ\text{C}$ · aluminum: $40.0 \times 0.897 = 35.88 \text{ J/}^\circ\text{C}$ · 1500. J each

$$\text{iron: } \Delta T = \frac{1500. \cancel{\text{J}}}{62.86 \cancel{\text{J/}^\circ\text{C}}} = +23.9 \text{ }^\circ\text{C} \rightarrow 30.0 \text{ }^\circ\text{C} + 23.9 \text{ }^\circ\text{C} = 53.9 \text{ }^\circ\text{C}$$

$$\text{aluminum: } \Delta T = \frac{1500. \cancel{\text{J}}}{35.88 \cancel{\text{J/}^\circ\text{C}}} = +41.8 \text{ }^\circ\text{C} \rightarrow 22.0 \text{ }^\circ\text{C} + 41.8 \text{ }^\circ\text{C} = \mathbf{63.8 \text{ }^\circ\text{C} \cdot \text{answer D}}$$

A took equal heat as an equal rise: that holds only when $m \cdot c$ matches, and 62.86 is not 35.88. B compared c alone: iron's 140. g makes its $m \cdot c$ the larger, so it rises only 23.9 °C. C stopped at the rise: 41.8 °C is how far aluminum climbs from 22.0 °C, not where it ends.

 **Sense check:** Aluminum starts 8.0 °C cooler but climbs about 18 °C further, so it finishes about 10 °C ahead. ✓



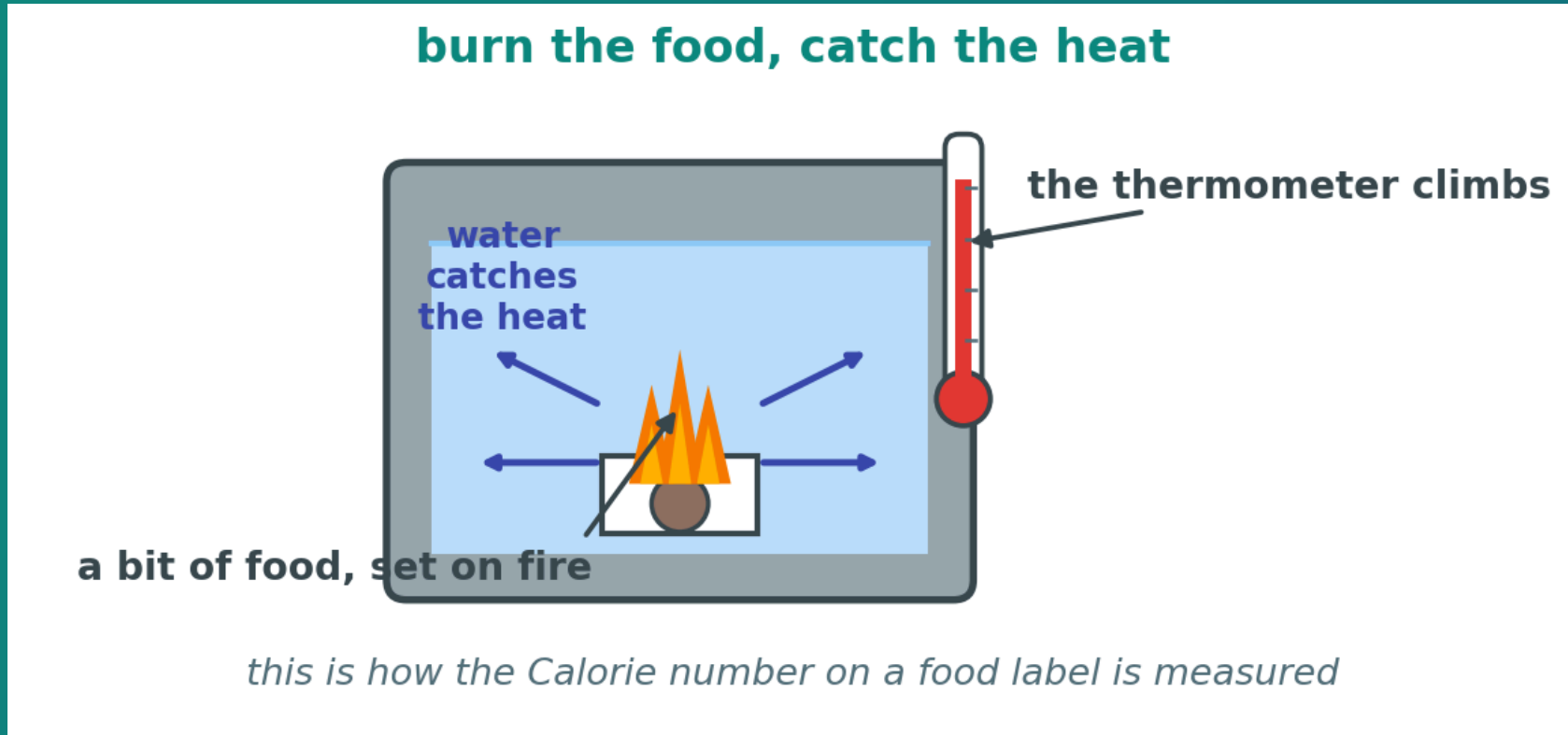
Check yourself

1. Equal masses of copper and water absorb the same heat from the same start. Which ends hotter, and which term in $\Delta T = q/mc$ decides it?
2. Two iron bars, 50 g and 150 g, absorb the same heat. Which shows the larger temperature change, and why?

5 • Calorimetry

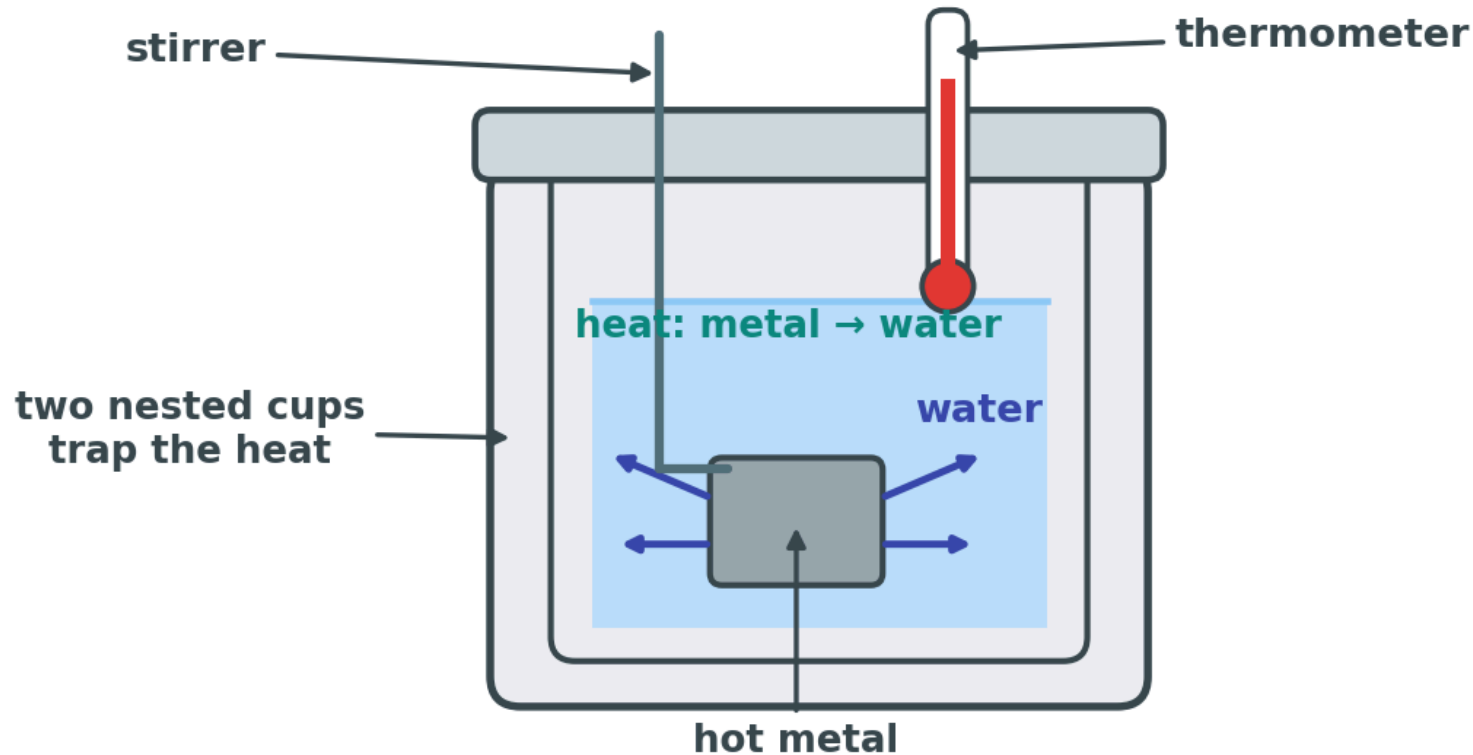
Use an insulated cup to make energy conservation visible (the heat one body loses equals the heat another gains) and solve for the heat a metal or a reaction in solution gives off, a final temperature, an unknown specific heat, or a food's Calories per gram, with any final temperature landing between the two starting temperatures.

How a Calorie gets measured



A food label's Calorie number is not estimated. The food is burned in a sealed chamber, water around it absorbs the heat, and the temperature rise gives the count.

Heat lost equals heat gained



Nest two foam cups, add a lid and a thermometer, and almost no heat escapes. Whatever heat the hot object loses, the water gains. The two settle at one shared temperature.

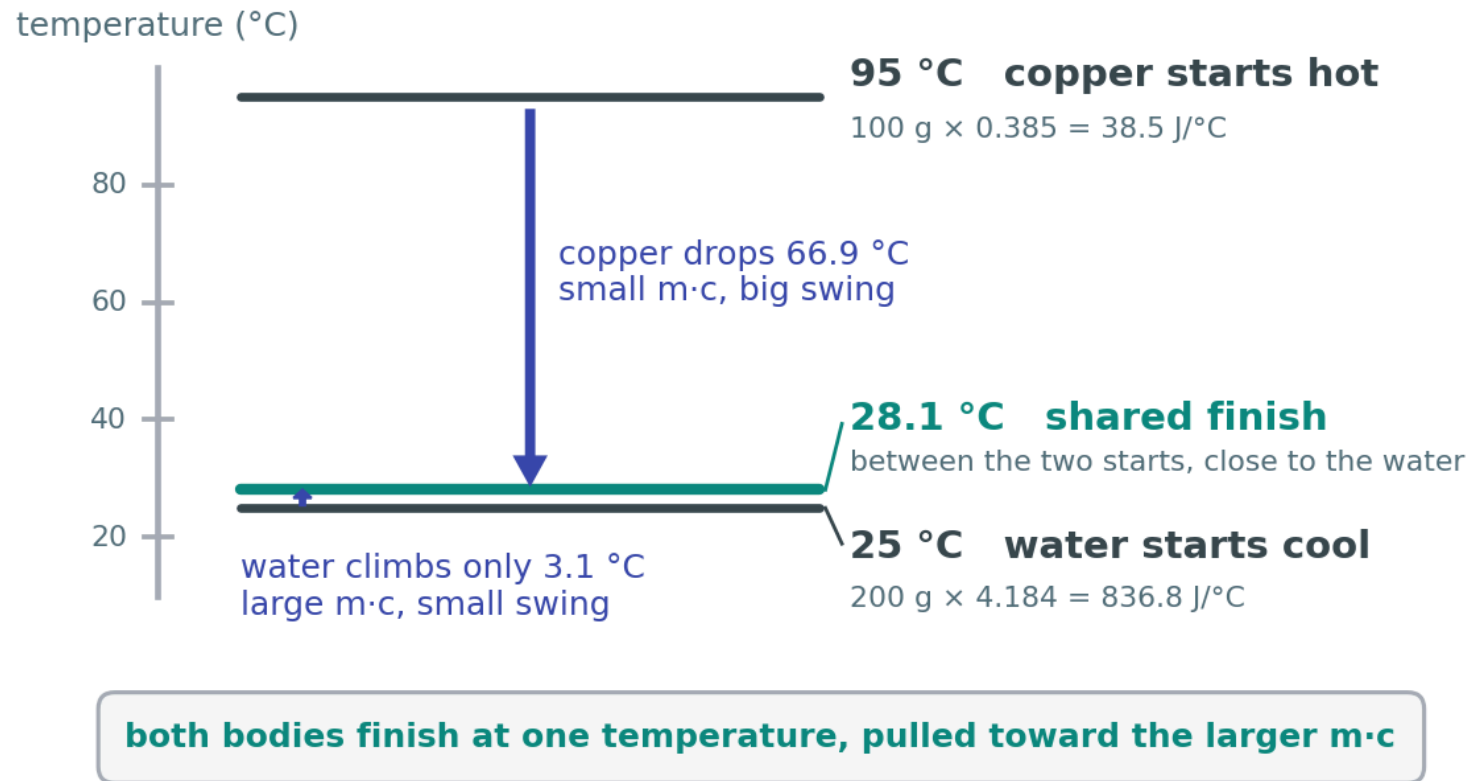
Energy is conserved in the cup

Inside the insulated cup, the energy that leaves one body enters the other. Energy is conserved: none is created or destroyed.

$$q_A + q_B = 0 \rightarrow q_A = -q_B$$

each body: $q = m \cdot c \cdot \Delta T$ with $\Delta T = T_f - T_i$ · one q comes out negative: that body released the heat

The final temperature sits between



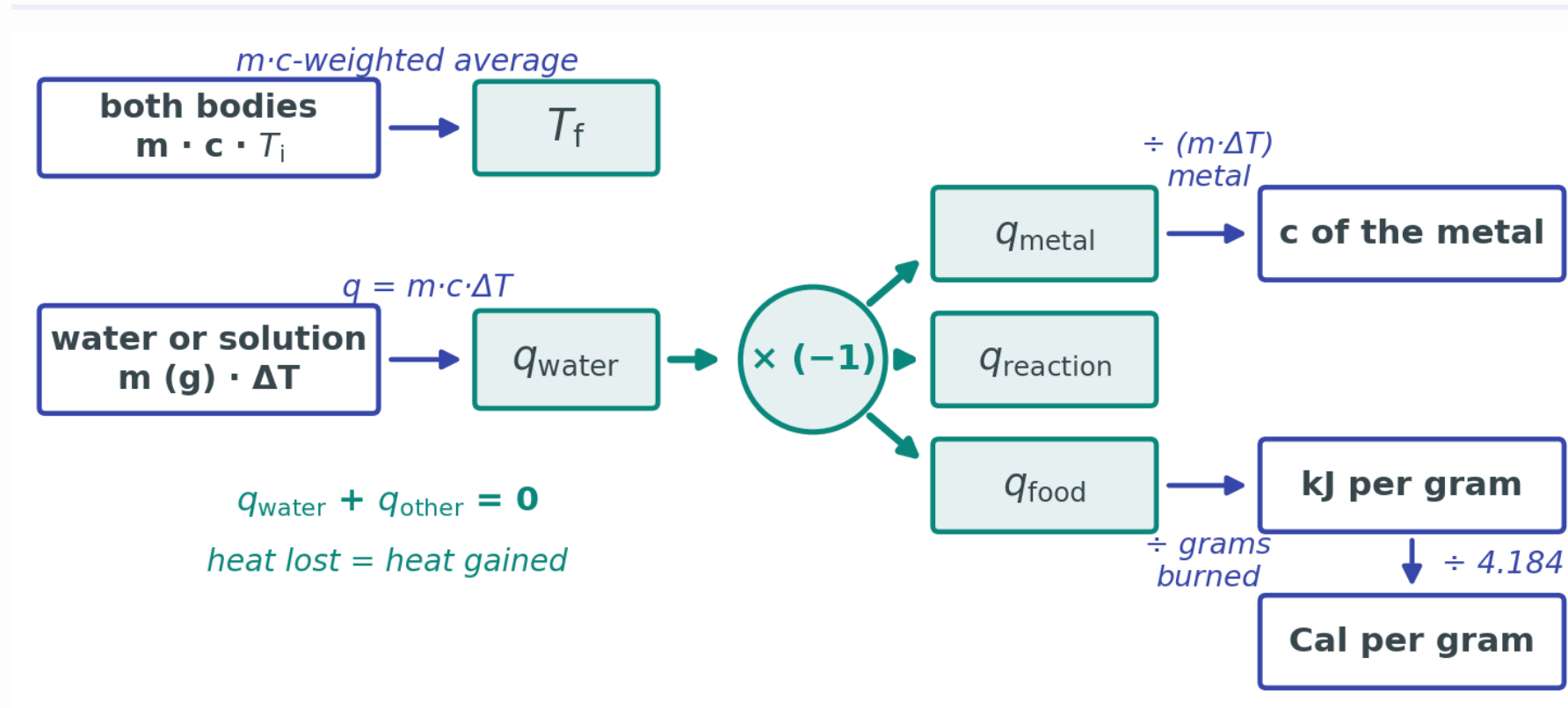
Both bodies end at the same temperature. It lands between the two starting temperatures, pulled toward whichever body carries the larger m·c: usually the water.



The method

1. **Set the balance:** $q_A = -q_B$, $\Delta T = T_f - T_i$ each.
2. **Solve without expanding:** T_f = the $m \cdot c$ -weighted average.
3. **Check:** T_f between the starts; heat lost = heat gained.

Every route starts from the water



The water's mass and ΔT give q_{water} . One sign flip gives the heat of the metal, reaction, or burning food in it. Two mixed solutions: add their volumes for the mass.



Guided example: a hot copper pipe

$$q_{\text{metal}} + q_{\text{water}} = 0 \rightarrow q_{\text{metal}} = -q_{\text{water}}$$

given: 72.5 g copper pipe, heated · 150.0 g water at 21.0 °C · final 33.6 °C · wanted: q of the pipe

A heated **72.5 g** copper pipe goes into a calorimeter holding **150.0 g** of water at 21.0 °C. Both settle at 33.6 °C. Find q for the pipe. Did the pipe gain or lose heat?

Start with the water: its mass and both temperatures are known.

✓ Guided example: solution

$$q_{\text{metal}} = -q_{\text{water}}$$

150.0 g water · $\Delta T = 33.6 - 21.0 = +12.6\text{ }^{\circ}\text{C}$ · the pipe's starting temperature was never measured

Step 1 · Set the balance

Heat gained by the water

$$q_{\text{water}} = 150.0\text{ g} \times 4.184\text{ J/g}\cdot^{\circ}\text{C} \times 12.6\text{ }^{\circ}\text{C} = +7.91 \times 10^3\text{ J}$$

Flip the sign

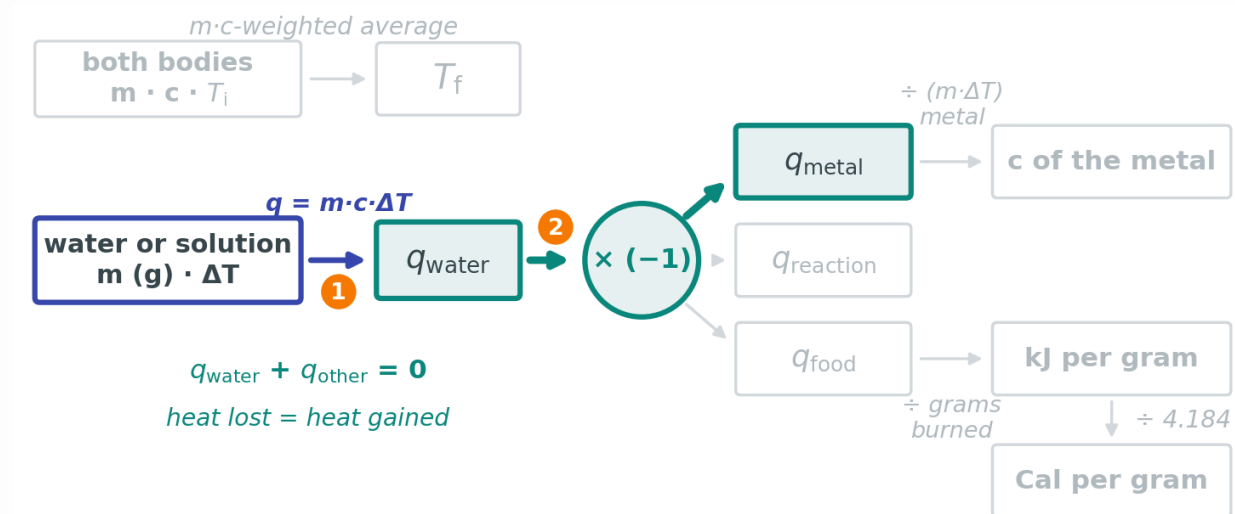
$$q_{\text{pipe}} = -q_{\text{water}} = -7.91 \times 10^3\text{ J}$$

 **Sense check:** The water warmed, so the pipe lost heat: its q is negative. The pipe's 72.5 g never entered; the water alone measured the heat. ✓

✓ Guided example: the route on the map

$$q_{\text{metal}} = -q_{\text{water}}$$

given: 150.0 g water, 21.0 → 33.6 °C · found: $q_{\text{pipe}} = -7.91 \times 10^3 \text{ J}$



🧠 **Sense check:** Two moves: $q = m \cdot c \cdot \Delta T$ for the water, then the sign flip. ✓



Worked example 1: mixing hot and cold water

$$q_{\text{hot}} = -q_{\text{cold}} \cdot \Delta T = T_f - T_i \text{ on both sides}$$

given: 150 g water at 70 °C · 100 g water at 20 °C · wanted: T_f

Pour **150 g** of water at 70 °C into **100 g** at 20 °C. Both are water, so both specific heats are 4.184. Find the final temperature.

✓ Worked example 1: solution

$$q_{\text{hot}} = -q_{\text{cold}} \cdot \Delta T = T_f - T_i \text{ on both sides}$$

150 g water at 70 °C · 100 g water at 20 °C · same $c = 4.184$, cancels

Step 1 · Set the balance

Step 2 · Solve without expanding

$$T_f = \frac{150 \times 70 + 100 \times 20}{150 + 100} = \frac{12,500}{250} = 50.0 \text{ °C}$$

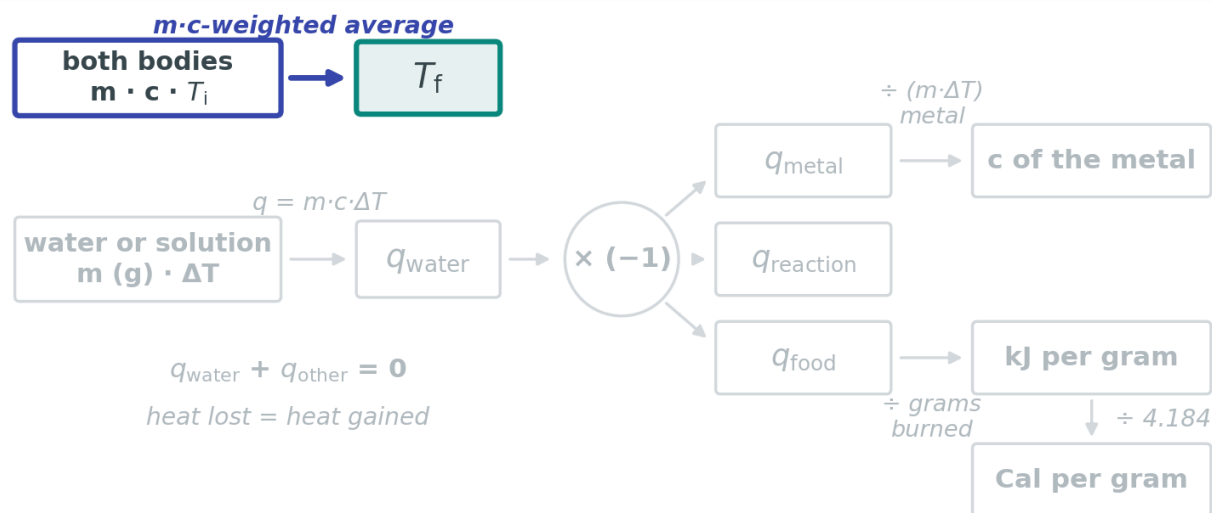
Step 3 · Check

🧠 **Sense check:** 50.0 °C lands between 20 and 70, closer to 70 because the hotter sample is heavier. And the books balance: the hot water released $150 \times 4.184 \times 20.0 = 12,552 \text{ J}$, the cold water absorbed $100 \times 4.184 \times 30.0 = 12,552 \text{ J}$. Heat lost = heat gained. ✓

✓ Worked example 1: the route on the map

$$q_{\text{hot}} = -q_{\text{cold}}$$

given: 150 g water at 70 °C · 100 g water at 20 °C · found: $T_f = 50.0$ °C



🧠 **Sense check:** With T_f unknown, neither q can come first. One move, the $m \cdot c$ -weighted average, gives T_f directly. ✓



Worked example 2: a hot bolt in water

$$q_{\text{iron}} = -q_{\text{water}} \cdot \Delta T = T_f - T_i \text{ on both sides}$$

given: 100 g iron at 90 °C · 150 g water at 20 °C · c: iron 0.449, water 4.184 · wanted: T_f

A **100 g iron** bolt at 90 °C drops into **150 g** of water at 20 °C.

A common first attempt: add the two heats. Test the result against the thermometer.

✓ Worked example 2: solution

$$q_{\text{iron}} = -q_{\text{water}} \cdot \Delta T = T_f - T_i \text{ on both sides}$$

100 g iron at 90 °C · 150 g water at 20 °C

A common first attempt

adding the heats: $T_f = 14.6 \text{ °C} \times$
 14.6 °C is below both 20 °C and 90 °C: no shared temperature sits outside the pair

Step 1 · Set the balance

$$44.9 \times (T_f - 90) = -627.6 \times (T_f - 20)$$

 **Sense check:** The two m·c terms are set: iron's $100 \times 0.449 = 44.9$ against the water's $150 \times 4.184 = 627.6$, each with its own $T_f - T_i$. The iron's ΔT will come out negative. ✓

✓ Worked example 2: the final temperature

$$q_{\text{iron}} = -q_{\text{water}} \cdot \Delta T = T_f - T_i \text{ on both sides}$$

m·c: iron 44.9, water 627.6 · start 90 °C and 20 °C

Step 2 · Solve without expanding

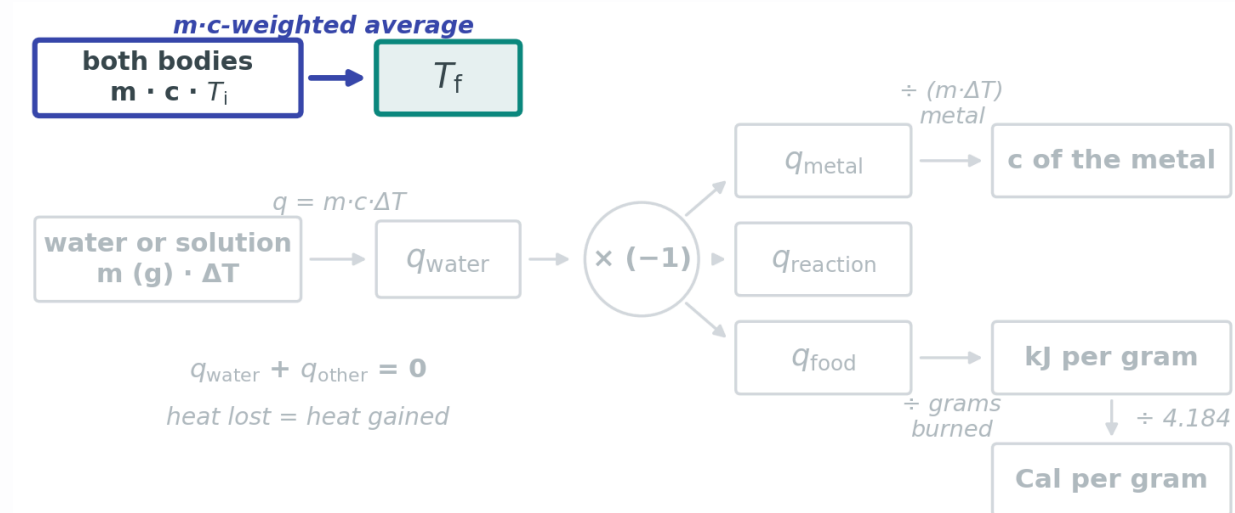
$$T_f = \frac{44.9 \times 90 + 627.6 \times 20}{44.9 + 627.6} = 24.7 \text{ °C}$$

🧠 **Sense check:** 24.7 °C sits between 20 and 90, close to the water: its m·c of 627.6 dwarfs the iron's 44.9. ✓

✓ Worked example 2: the route on the map

$$q_{\text{iron}} = -q_{\text{water}}$$

given: 100 g iron at 90 °C · 150 g water at 20 °C · found: $T_f = 24.7$ °C



🧠 **Sense check:** The same one move as two water samples. Only the m·c weights changed. ✓

Take-home: one loses, the other gains

heat lost by the hot body = heat gained by the cold body

T_f always lands between the two starting temperatures

Never add the two heats. One body releases, the other absorbs. Set the loss equal to the gain, and a final temperature outside the starting pair means a dropped sign.

Your turn: aluminum into water

Drop **80 g aluminum** ($c = 0.897$) at $100\text{ }^{\circ}\text{C}$ into **200 g** water at $22\text{ }^{\circ}\text{C}$. Fill the m·c terms into the weighted average that solves $q_{\text{Al}} = -q_{\text{water}}$.

$$T_f = \frac{\underline{\hspace{1cm}} \times 100 + \underline{\hspace{1cm}} \times 22}{\underline{\hspace{1cm}} + \underline{\hspace{1cm}}} = \underline{\hspace{1cm}}\text{ }^{\circ}\text{C}$$

$$T_f = \frac{71.76 \times 100 + 836.8 \times 22}{71.76 + 836.8} = \mathbf{28.2\text{ }^{\circ}\text{C}}$$

 **Sense check:** Between 22 and 100, and close to 22: the water's m·c of 836.8 far outweighs the aluminum's 71.76.

✓

The bomb calorimeter: the water does the measuring

$$q_{\text{food}} = -q_{\text{water}} = m \cdot c \cdot \Delta T$$

$$\text{water's mass from its volume at } 1.00 \text{ g/mL} \cdot \text{kJ} \div \text{grams burned} = \text{kJ/g} \cdot 1 \text{ Cal} = 4.184 \text{ kJ}$$

Food burns in a sealed steel chamber sunk in water. The water absorbs all the heat, so its temperature rise measures it. Divide by the grams burned for the energy per gram.



Worked example 3: Calories per gram of a candy bar

$$q_{\text{food}} = -q_{\text{water}} \cdot q_{\text{water}} = m \cdot c \cdot \Delta T$$

given: 5.00 g of candy bar burned · 1840 mL water · 20.0 → 35.0 °C · wanted: heat released per gram,
in kJ/g and Cal/g

A **5.00 g** piece of a candy bar burns in a bomb calorimeter holding **1840 mL** of water. The water warms from 20.0 to 35.0 °C. Find the heat released per gram, in kJ/g and in Cal/g.

✓ Worked example 3: solution

$$q_{\text{food}} = -q_{\text{water}} \cdot q_{\text{water}} = m \cdot c \cdot \Delta T$$

$$5.00 \text{ g burned} \cdot 1840 \text{ mL water} = 1840 \text{ g at } 1.00 \text{ g/mL} \cdot \Delta T = 35.0 - 20.0 = +15.0 \text{ }^{\circ}\text{C}$$

Two moves are needed.

Step 1 • Heat gained by the water

$$q_{\text{water}} = 1840 \text{ g} \times 4.184 \text{ J/g} \cdot ^{\circ}\text{C} \times 15.0 \text{ }^{\circ}\text{C} = 115,478 \text{ J} = 1.15 \times 10^5 \text{ J}$$

Step 2 • Per gram, then Calories

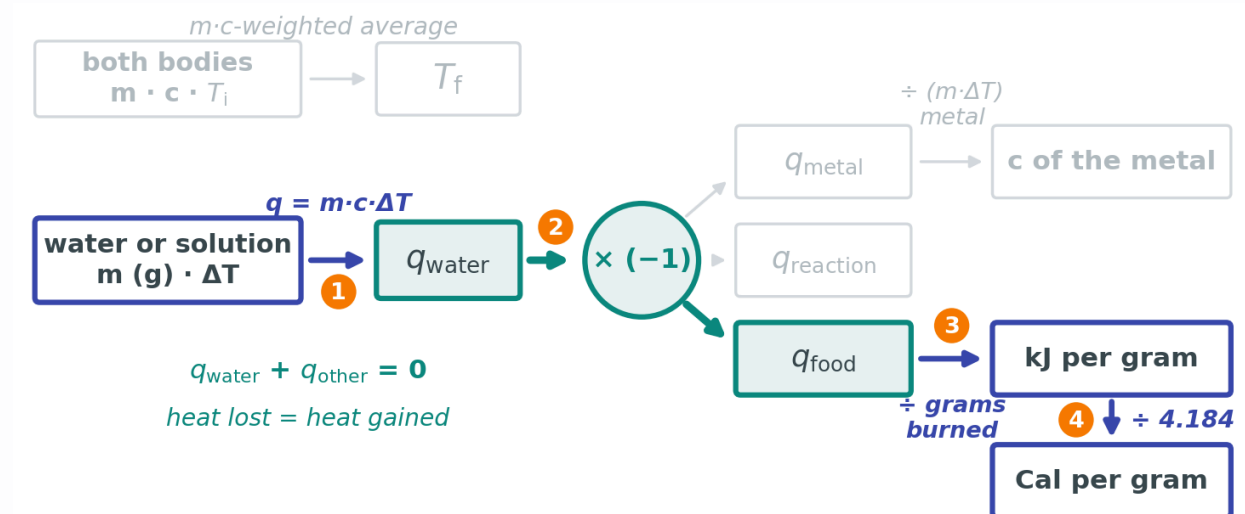
$$\frac{115.5 \text{ kJ}}{5.00 \text{ g}} = 23.1 \text{ kJ/g} \rightarrow 23.1 \text{ kJ/g} \times \frac{1 \text{ Cal}}{4.184 \text{ kJ}} = 5.52 \text{ Cal/g}$$

 **Sense check:** A 40.0 g bar of it releases $40.0 \text{ g} \times 5.52 \text{ Cal/g} = 221 \text{ Cal}$. Fat carries about 9 Cal/g and sugar about 4, so a bar of both lands near 5.5. ✓

✓ Worked example 3: the route on the map

$$q_{\text{food}} = -q_{\text{water}}$$

given: 5.00 g burned · 1840 g water, 20.0 → 35.0 °C · found: 23.1 kJ/g = 5.52 Cal/g



🧠 **Sense check:** Step 1 of the solution is arrows 1 and 2: the water's q , then the sign flip. Step 2 is arrows 3 and 4. ✓

⚠ Where this goes wrong

100 g iron at 90 °C into 150 g water at 20 °C → $T_f = 24.7$ °C
the correct final temperature both bodies reach

Ignoring the water's share. Assuming the metal just cools to the water's temperature gives $T_f = 20$ °C. The water warms too; its gain is part of the balance.

Adding the two heats. Same sign on both sides gives $T_f = 14.6$ °C, below both starting temperatures. One body loses heat, the other gains it: heat lost = heat gained.

Swapping the masses onto the wrong specific heats. Pairing 150 with the iron and 100 with the water gives $T_f = 29.7$ °C. Each mass keeps its own substance's c .



Practice 1

$$q_{\text{metal}} = -q_{\text{water}}$$

given: 45.0 g iron bearing, hot · 110.0 g water at 19.5 °C · both end at 30.8 °C · c: iron 0.449, water 4.184 · wanted: q of the bearing

A **45.0 g** iron bearing, hot from a grinder, drops into **110.0 g** of water at 19.5 °C. The water warms to 30.8 °C. What is q for the bearing, in joules?

- A. -2.13×10^3
- B. 5.20×10^3
- C. -558
- D. -5.20×10^3

✓ Practice 1 · answer: D

$$q_{\text{metal}} = -q_{\text{water}}$$

110.0 g water · $\Delta T = 30.8 - 19.5 = +11.3\text{ }^{\circ}\text{C}$ · the bearing's own ΔT is unknown

$$q_{\text{water}} = 110.0\text{ g} \times 4.184\text{ J/g}\cdot^{\circ}\text{C} \times 11.3\text{ }^{\circ}\text{C} = +5.20 \times 10^3\text{ J} \rightarrow$$

$$q_{\text{bearing}} = -5.20 \times 10^3\text{ J (answer D)}$$

A put the bearing's 45.0 g into the water's q : $-(45.0 \times 4.184 \times 11.3) = -2.13 \times 10^3$. B stopped at q_{water} : $+5.20 \times 10^3$ is the heat the water gained, before the sign flip. C used iron's specific heat for the water: $-(110.0 \times 0.449 \times 11.3) = -558$.

🧠 **Sense check:** The water warmed, so the bearing lost heat: q is negative, and it matches the water's gain in size. ✓



Practice 2

$$q_{\text{reaction}} = -q_{\text{solution}}$$

given: 60.0 mL acid + 60.0 mL base, both at 20.8 °C · peak 33.9 °C · treat as water: 1.00 g/mL, $c = 4.184$

In a coffee-cup calorimeter, **60.0 mL** of hydrochloric acid and **60.0 mL** of sodium hydroxide solution, both at 20.8 °C, are mixed. The temperature peaks at 33.9 °C. What is q for the reaction, in joules?

- A. -6.58×10^3
- B. -3.29×10^3
- C. 6.58×10^3
- D. -1.70×10^4

✓ Practice 2 · answer: A

$$q_{\text{reaction}} = -q_{\text{solution}}$$

$$60.0 \text{ mL} + 60.0 \text{ mL} = 120.0 \text{ mL} = 120.0 \text{ g at } 1.00 \text{ g/mL} \cdot \Delta T = 33.9 - 20.8 = +13.1 \text{ }^{\circ}\text{C}$$

$$q_{\text{solution}} = 120.0 \text{ g} \times 4.184 \text{ J/g}\cdot^{\circ}\text{C} \times 13.1 \text{ }^{\circ}\text{C} = +6.58 \times 10^3 \text{ J} \rightarrow$$

$$q_{\text{reaction}} = \mathbf{-6.58 \times 10^3 \text{ J (answer A)}}$$

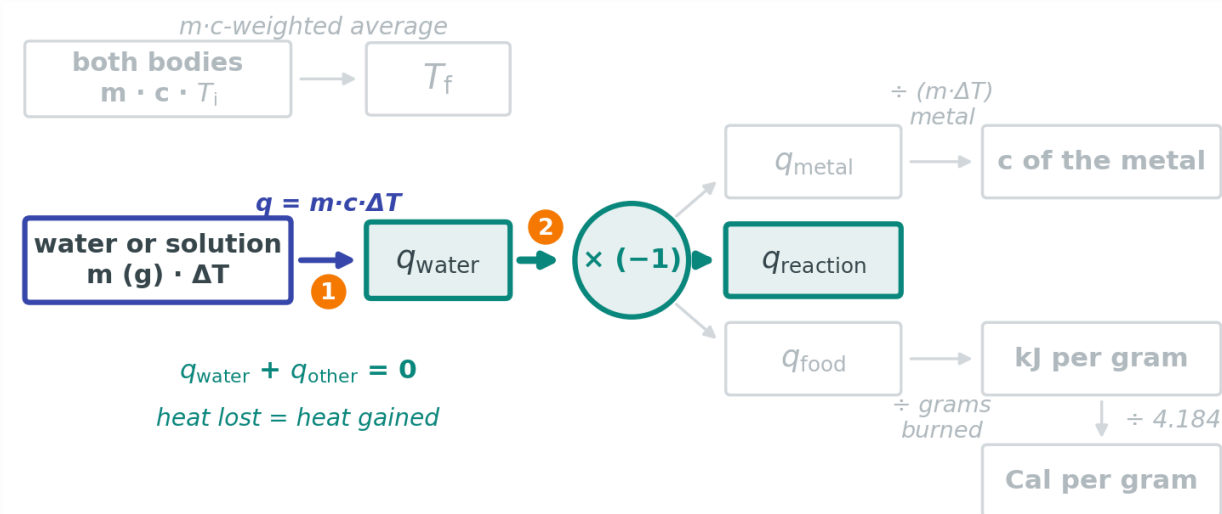
B counted one solution's 60.0 g: the whole 120.0 g mixture warmed, so $-(60.0 \times 4.184 \times 13.1) = -3.29 \times 10^3$ is half the heat. C stopped at q_{solution} : $+6.58 \times 10^3$, before the sign flip. D used the peak temperature as ΔT : $-(120.0 \times 4.184 \times 33.9) = -1.70 \times 10^4$.

 **Sense check:** The mixture warmed, so the reaction released heat: q_{reaction} is negative, exothermic. ✓

✓ Practice 2: the route on the map

$$q_{\text{reaction}} = -q_{\text{solution}}$$

given: 60.0 mL + 60.0 mL mixed, 20.8 → 33.9 °C · found: $q_{\text{reaction}} = -6.58 \times 10^3 \text{ J}$



Sense check: The whole 120.0 g mixture is the water; flip its q for the reaction. ✓



Practice 3

heat lost = heat gained

given: 60.0 g copper at 120.0 °C · 125 g water at 25.0 °C · c: copper 0.385, water 4.184 · wanted: T_f

A **60.0 g copper** block at 120.0 °C drops into **125 g** of water at 25.0 °C. What final temperature, in °C, do they reach?

- A. 29.0
- B. 40.3
- C. 55.8
- D. 72.5

✓ Practice 3 · answer: A

heat lost = heat gained

m·c: copper $60.0 \times 0.385 = 23.1$ · water $125 \times 4.184 = 523$ · start 120.0 °C and 25.0 °C

$$T_f = \frac{23.1 \times 120.0 + 523 \times 25.0}{23.1 + 523} = 29.0\text{ °C (answer A)}$$

B swapped the masses onto the wrong specific heats: 40.3 °C . C dropped the specific heats and weighted by mass alone: $(60.0 \times 120.0 + 125 \times 25.0) \div 185 = 55.8\text{ °C}$, a shortcut that works only when both bodies are water. D took the plain average of the starts: $(120.0 + 25.0) \div 2 = 72.5\text{ °C}$, as if the two bodies weighed the same and shared one c.

 **Sense check:** 29.0 °C sits between 25 and 120, close to the water: its m·c of 523 outweighs the copper's 23.1. ✓



Practice 4

$$q_{\text{metal}} = -q_{\text{water}} \cdot \Delta T = T_f - T_i \text{ on both sides}$$

given: 95.0 g metal at 99.0 °C · 40.0 g water at 18.0 °C · measured: $T_f = 34.5$ °C · wanted: c of the metal

A **95.0 g** sample of an unknown metal at 99.0 °C drops into **40.0 g** of water at 18.0 °C. Both settle at **34.5 °C**. What is the metal's specific heat, in J/g·°C?

- A. -0.451
- B. 2.76×10^3
- C. 0.843
- D. 0.451
- E. 2.54

✓ Practice 4 · answer: D

$$q_{\text{metal}} = -q_{\text{water}} \cdot \Delta T = T_f - T_i \text{ on both sides}$$

$$\text{water: } 40.0 \text{ g} \cdot \Delta T = 34.5 - 18.0 = +16.5 \text{ }^{\circ}\text{C} \cdot \text{metal: } 95.0 \text{ g} \cdot \Delta T = 34.5 - 99.0 = -64.5 \text{ }^{\circ}\text{C}$$

$$q_{\text{water}} = 40.0 \text{ g} \times 4.184 \text{ J/g}\cdot^{\circ}\text{C} \times 16.5 \text{ }^{\circ}\text{C} = 2761 \text{ J} \rightarrow c = \frac{-2761 \text{ J}}{95.0 \text{ g} \times (-64.5 \text{ }^{\circ}\text{C})}$$

$$= 0.451 \text{ J/g}\cdot^{\circ}\text{C} \text{ (answer D)}$$

A dropped the minus in $q_{\text{metal}} = -q_{\text{water}}$: $2761 \div (95.0 \times (-64.5)) = -0.451$, a negative specific heat. B stopped at q_{water} : 2761 J, before dividing by the metal's $m \cdot \Delta T$. C used the final temperature as the metal's ΔT : $2761 \div (95.0 \times 34.5) = 0.843$. E swapped the two masses: $-(95.0 \times 4.184 \times 16.5) \div (40.0 \times (-64.5)) = 2.54$.

 **Sense check:** The metal fell 64.5 °C while the water rose only 16.5 °C: the metal's $m \cdot c$, $95.0 \times 0.451 = 42.8$, is about a quarter of the water's 167. A c near 0.45, like iron's, fits. ✓



Practice 5

$$q_{\text{food}} = -q_{\text{water}} \cdot q_{\text{water}} = m \cdot c \cdot \Delta T$$

given: 52.7 g granola bar · 4.00 g of it burned · 1650 mL water · 20.5 → 32.1 °C · wanted: Calories released by the whole bar

A food lab burns a **4.00 g** chip of a **52.7 g** granola bar in a bomb calorimeter. The **1650 mL** water jacket climbs from 20.5 to 32.1 °C. How many Calories does the whole bar release?

- A. 252
- B. 4.41×10^3
- C. 19.1
- D. 698
- E. 1.06×10^3

✓ Practice 5 · answer: A

$$q_{\text{food}} = -q_{\text{water}} \cdot q_{\text{water}} = m \cdot c \cdot \Delta T$$

$$1650 \text{ mL water} = 1650 \text{ g} \cdot \Delta T = 32.1 - 20.5 = +11.6 \text{ }^{\circ}\text{C} \cdot 4.00 \text{ g burned} \cdot \text{whole bar } 52.7 \text{ g}$$

$$q_{\text{water}} = 1650 \text{ g} \times 4.184 \text{ J/g}^{\circ}\text{C} \times 11.6 \text{ }^{\circ}\text{C} = 80,082 \text{ J} \rightarrow \frac{80.08 \text{ kJ}}{4.00 \text{ g}} = 20.0 \text{ kJ/g}$$

$$52.7 \text{ g} \times \frac{20.0 \text{ kJ}}{1 \text{ g}} \times \frac{1 \text{ Cal}}{4.184 \text{ kJ}} = \mathbf{252 \text{ Cal (answer A)}}$$

B multiplied by 4.184: $1055 \times 4.184 = 4.41 \times 10^3$. C skipped the scale-up: $80.08 \text{ kJ} \div 4.184 = 19.1$, the chip alone. D used $32.1 \text{ }^{\circ}\text{C}$ as ΔT : 698. E stopped at kilojoules: 1.06×10^3 .

 **Sense check:** 252 Cal for a 52.7 g bar is about 4.8 Cal/g, between sugar's 4 and fat's 9, where a granola bar sits. ✓



Check yourself

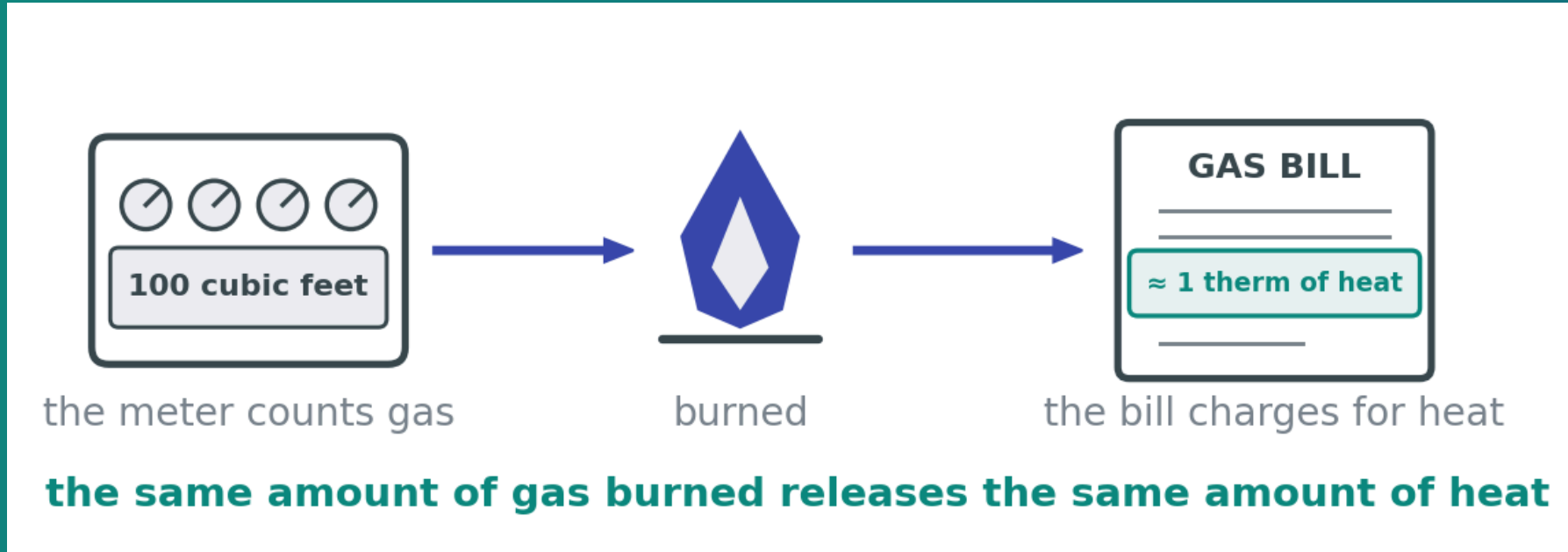
1. A hot metal is dropped into cool water in an insulated cup. Write the conservation law that relates the heat the metal loses to the heat the water gains.
2. A final temperature comes out below both starting temperatures. What went wrong, and what does the correct answer always sit between?

A reaction releases a fixed amount of heat per mole, and the calorimeter is how that number gets measured. Written beside its balanced equation as ΔH , it becomes a conversion factor as real as any molar mass.

6 • Thermochemical Equations

Treat ΔH as tied to the equation as written: reverse the equation and flip its sign, scale the coefficients and scale ΔH in step, and use ΔH as a conversion factor to find the heat released or absorbed by a given amount of substance.

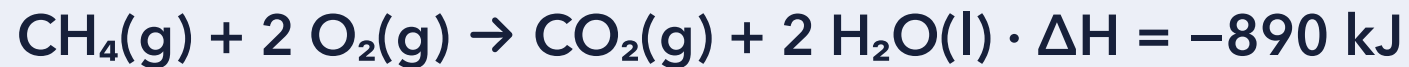
The gas bill charges for heat



The utility charges for energy delivered, not the gas itself. Burning a set amount of fuel releases a set amount of heat, and the bill counts that heat.

ΔH is tied to the equation as written

A thermochemical equation pairs a balanced equation with its ΔH . That ΔH belongs to those exact coefficients and states. Change the equation, and ΔH changes with it.



–890 kJ is released when 1 mol CH_4 burns exactly as written

Reverse the equation, flip the sign

Running a reaction backward reverses its heat flow. A release becomes an equal absorption. The magnitude is unchanged; only the sign flips.



forward, $\Delta\text{H} = -286 \text{ kJ}$ (released) · reversed, $\Delta\text{H} = +286 \text{ kJ}$ (absorbed)

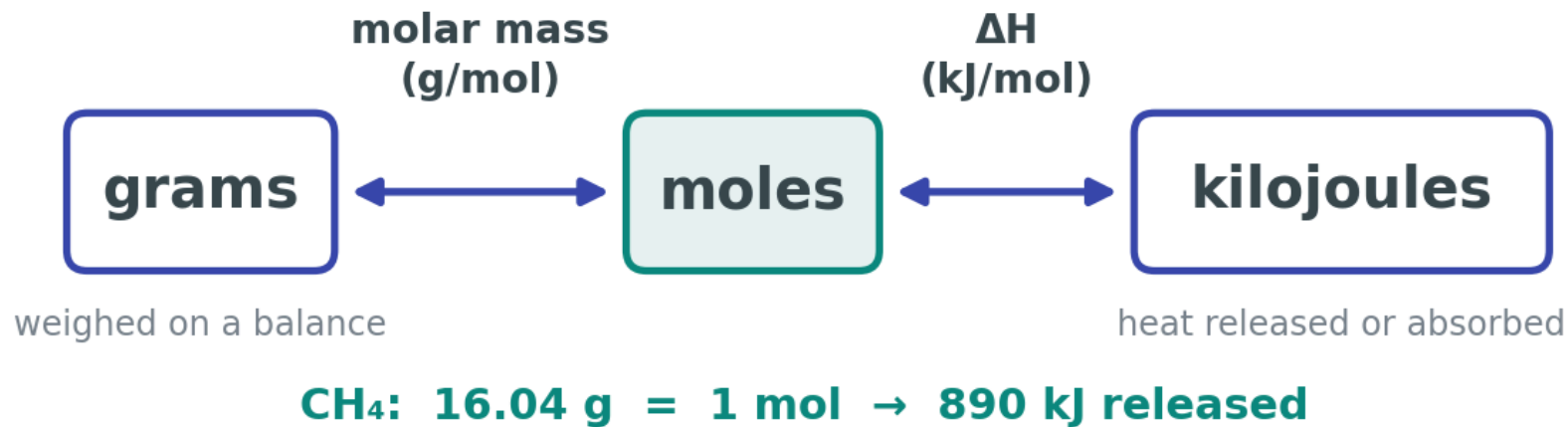
Scale the coefficients, scale ΔH

ΔH is proportional to the amount reacting. Double every coefficient and twice as much reacts, so ΔH doubles. Halve them and ΔH halves.



from $\text{H}_2 + \frac{1}{2} \text{O}_2 \rightarrow \text{H}_2\text{O}$, $\Delta H = -286 \text{ kJ}$, scaled $\times 2$

ΔH is a conversion factor



The coefficients turn ΔH into a factor: kJ per mole of any species in the equation. Chain it with molar mass to move between grams of fuel and kilojoules of heat.



The method

1. **Match the target to the given.** Reverse the equation if the target runs backward.
2. **Flip the sign** for a reversal; **scale ΔH** by the factor that scales the coefficients.
3. **For heat,** use ΔH as kJ per mole and cancel units.



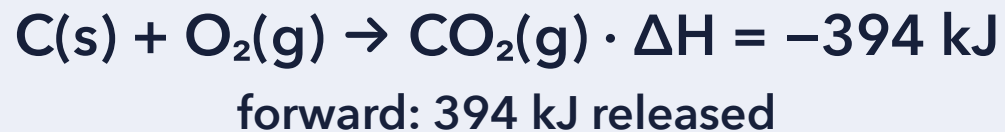
Worked example 1: reversing an equation



given: this equation and ΔH · wanted: ΔH for the reverse

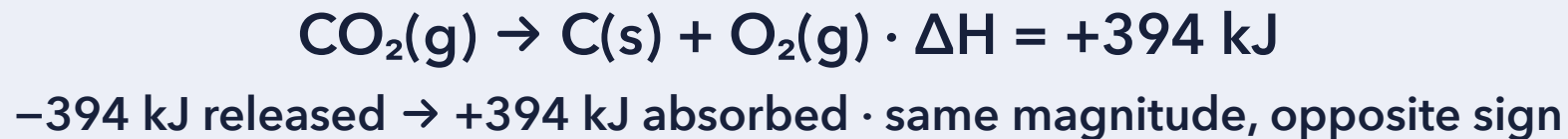
Carbon burns to carbon dioxide, releasing 394 kJ. What is ΔH for the reverse, $\text{CO}_2\text{(g)} \rightarrow \text{C(s)} + \text{O}_2\text{(g)}$?

✓ Worked example 1: solution



Step 1 · Match the target to the given

Step 2 · Flip the sign



🧠 **Sense check:** Splitting CO₂ must cost exactly the energy its formation released. ✓



Worked example 2: reverse and scale



given: this equation and ΔH · wanted: ΔH for $2 \text{H}_2\text{O}(\text{l}) \rightarrow 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g})$

Find ΔH for $2 \text{H}_2\text{O}(\text{l}) \rightarrow 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g})$.

A common first attempt: flip the sign but leave the coefficients' change out of ΔH . Test it.

✓ Worked example 2: solution



Step 1 · Match the target to the given

Step 2 · Flip the sign, then scale ΔH

$$\Delta\text{H} = -(2 \times -286 \text{ kJ}) = +572 \text{ kJ}$$

 **Sense check:** Flip only: +286 kJ ✗. Scale only: -572 kJ ✗. Both moves give +572 kJ: forming 2 mol water released 572 kJ, so splitting it absorbs 572 kJ. ✓

Take-home: reverse flips, scale multiplies

flip the arrow, flip the sign · scale the equation, scale the ΔH

reversed: $\Delta H \times (-1)$ · scaled by n: $\Delta H \times n$ · do both when both apply

Each change to the equation changes ΔH . Reversing flips the sign. Scaling multiplies. A target that is both reversed and doubled needs both moves. Say the rule until it is automatic.

Your turn: reverse and scale

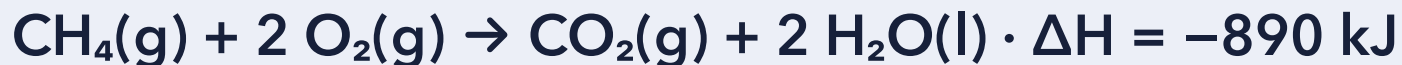
Given $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{SO}_3(\text{g})$, $\Delta H = -198.2 \text{ kJ}$, find ΔH for $4 \text{SO}_3(\text{g}) \rightarrow 4 \text{SO}_2(\text{g}) + 2 \text{O}_2(\text{g})$ (reversed and doubled).

$$\Delta H = \underline{\hspace{1cm}} (\underline{\hspace{1cm}} \times -198.2 \text{ kJ}) = \underline{\hspace{1cm}} \text{ kJ}$$

$$\Delta H = -(2 \times -198.2 \text{ kJ}) = \mathbf{+396.4 \text{ kJ}}$$

 **Sense check:** Reversed, so the sign flips; doubled, so $\times 2$. Forming 4 mol SO_3 released 396.4 kJ, so the reverse absorbs it. ✓

Reading ΔH per mole of each substance



-890 kJ goes with 1 mol CH_4 · with 2 mol O_2 · with 1 mol CO_2 · with 2 mol H_2O

Pick one substance. Its coefficient, in moles, goes under ΔH , either way up.

$$\frac{-890 \text{ kJ}}{1 \text{ mol CH}_4} \text{ or } \frac{1 \text{ mol CH}_4}{-890 \text{ kJ}} \cdot \frac{-890 \text{ kJ}}{2 \text{ mol O}_2} \text{ or } \frac{2 \text{ mol O}_2}{-890 \text{ kJ}}$$

Subscripts never enter the factor. The 2 in O_2 counts atoms inside one molecule.

Guided example: a camping stove



given: 0.250 mol C_3H_8 · wanted: ΔH in kJ

A camping stove burns **0.250 mol** of propane. What is ΔH for this amount?

Read the equation first: burning 1 mol C_3H_8 releases 2220 kJ.

✓ Guided example: solution



given: 0.250 mol C_3H_8 · -2220 kJ per 1 mol C_3H_8

Step 1 · Write ΔH per mole of C_3H_8

Step 2 · Cancel mol C_3H_8

$$0.250 \text{ mol } \cancel{\text{C}_3\text{H}_8} \times \frac{-2220 \text{ kJ}}{1 \cancel{\text{ mol } \text{C}_3\text{H}_8}} = -555 \text{ kJ}$$

 **Sense check:** A quarter of 2220 kJ is 555 kJ, released, so ΔH is negative. The conversion factor and the recipe agree. ✓

✓ Guided example: counted from the oxygen



0.250 mol C₃H₈ burns with $0.250 \times 5 = 1.25 \text{ mol O}_2 \cdot -2220 \text{ kJ per } 5 \text{ mol O}_2$

Write ΔH per mole of O₂

Cancel mol O₂

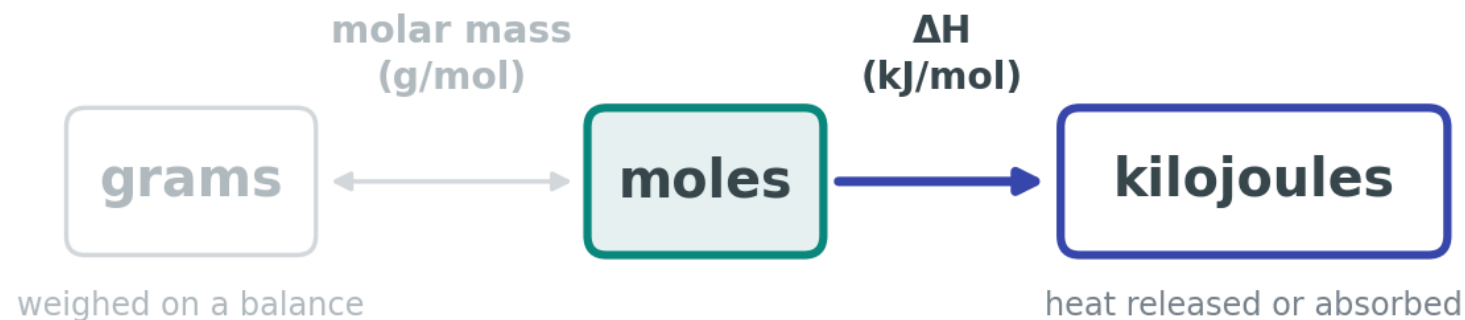
$$1.25 \cancel{\text{ mol O}_2} \times \frac{-2220 \text{ kJ}}{5 \cancel{\text{ mol O}_2}} = -555 \text{ kJ}$$

 **Sense check:** Same heat as the propane count. Keeping $-2220 \text{ kJ per } 1 \text{ mol O}_2$ gives $1.25 \times -2220 = -2775 \text{ kJ}$, five times too much. ✓

✓ Guided example: the route on the map



given: 0.250 mol C_3H_8 · found: -555 kJ



C_3H_8 : 0.250 mol → 555 kJ released

🧠 **Sense check:** Moles in hand: one arrow, one conversion factor, ΔH per mole of C_3H_8 . The grams box is not needed. ✓



Practice 1: moles of I₂ to kilojoules



given: 2.50 mol I₂ · wanted: ΔH in kJ

What is ΔH, in kJ, when **2.50 mol** of I₂ reacts?

- A. +66.3
- B. −133
- C. +133
- D. +0.0472

✓ Practice 1 · answer: C



2.50 mol I_2 · +53.0 kJ per 1 mol I_2

$$2.50 \cancel{\text{mol I}_2} \times \frac{+53.0 \text{ kJ}}{1 \cancel{\text{mol I}_2}} = +133 \text{ kJ (answer C)}$$

A put HI's coefficient under ΔH : $2.50 \times 53.0 \div 2 = +66.3 \text{ kJ}$, but I_2 's coefficient is 1. B gave the sign of a release: ΔH is positive, so heat is absorbed. D flipped the factor: $2.50 \div 53.0 = 0.0472$, and mol never cancels.

 **Sense check:** Each mole of I_2 absorbs 53.0 kJ, and 2.50 mol absorbs 2.5 times that: about 130 kJ. ✓



Practice 2: moles of NH_3 to kilojoules



given: 3.60 mol NH_3 formed · wanted: ΔH in kJ

What is ΔH , in kJ, when **3.60 mol** of NH_3 forms?

- A. -111
- B. -166
- C. -0.0781
- D. -332

✓ Practice 2 · answer: B



$$3.60 \text{ mol NH}_3 \cdot -92.2 \text{ kJ per 2 mol NH}_3$$

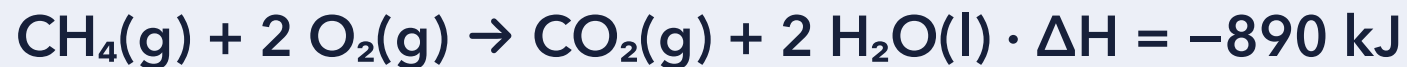
$$3.60 \text{ mol NH}_3 \times \frac{-92.2 \text{ kJ}}{2 \text{ mol NH}_3} = -166 \text{ kJ (answer B)}$$

D kept $-92.2 \text{ kJ per 1 mol NH}_3$: $3.60 \times -92.2 = -332 \text{ kJ}$, double the answer, because the equation forms 2 mol. A put H_2 's coefficient under ΔH : $3.60 \times -92.2 \div 3 = -111 \text{ kJ}$. C flipped the factor: $3.60 \times 2 \div -92.2 = -0.0781$, and mol never cancels.

 **Sense check:** Per mole of NH_3 the heat is $-92.2 \div 2 = -46.1 \text{ kJ}$, and $3.60 \text{ mol} \times 46 \text{ kJ}$ is near 166 kJ released. ✓



Worked example 3: grams of fuel to kilojoules

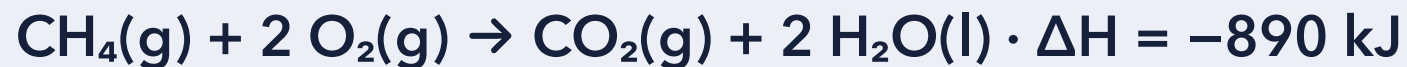


given: 40.0 g CH₄ burned · molar mass CH₄ = 16.04 g/mol · wanted: ΔH

Burn **40.0 g** of methane completely. What is ΔH for this amount of fuel?

Two conversion factors are needed: grams to moles, then moles to kilojoules.

✓ Worked example 3: solution



$$40.0 \text{ g CH}_4 \cdot 16.04 \text{ g/mol} \cdot -890 \text{ kJ per 1 mol CH}_4$$

Two conversion factors are needed.

Step 1 · Pick the ΔH orientation

Step 2 · Chain grams $\rightarrow$ moles $\rightarrow$ kilojoules

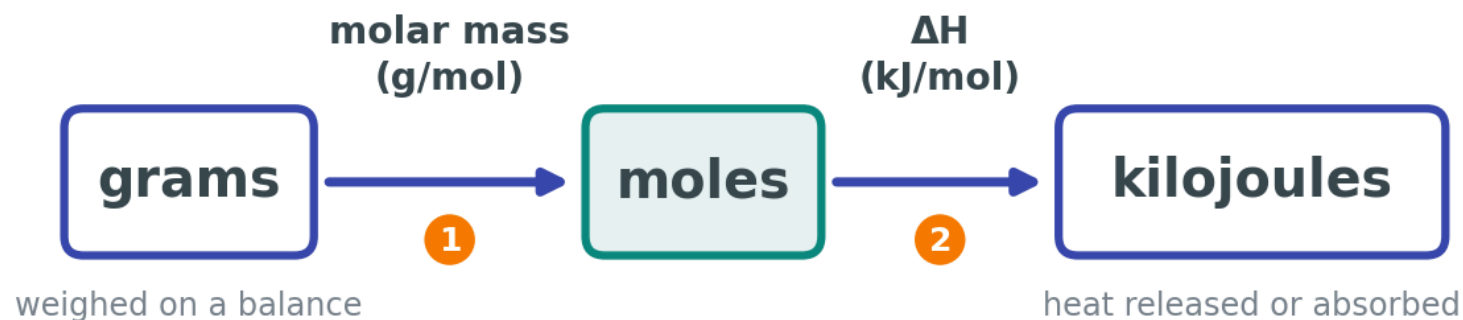
$$40.0 \text{ g} \times \frac{1 \text{ mol}}{16.04 \text{ g}} \times \frac{-890 \text{ kJ}}{1 \text{ mol}} = -2220 \text{ kJ}$$

🧠 **Sense check:** 40.0 g is about 2.5 mol, each releasing 890 kJ: $\Delta H \approx -2220 \text{ kJ}$, about 2200 kJ given off. ✓

✓ Worked example 3: the route on the map



given: 40.0 g CH₄ · found: -2220 kJ



CH₄: 40.0 g → 2.49 mol → 2220 kJ released

🧠 **Sense check:** Grams in hand: two arrows, two conversion factors. Molar mass first, then ΔH per mole of CH₄. ✓



Worked example 4: when the coefficient is 2

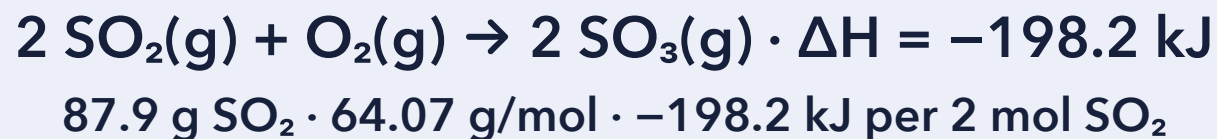


given: 87.9 g SO_2 · molar mass SO_2 = 64.07 g/mol · wanted: heat evolved in kJ

Calculate the heat evolved when **87.9 g of sulfur dioxide** is converted to sulfur trioxide.

Same grams → moles → kilojoules chain as the methane problem. SO_2 's coefficient is 2, and that 2 goes under ΔH .

✓ Worked example 4: solution



Step 1 • Write ΔH per mole of SO_2

Step 2 • Chain grams $\rightarrow$ moles $\rightarrow$ kilojoules

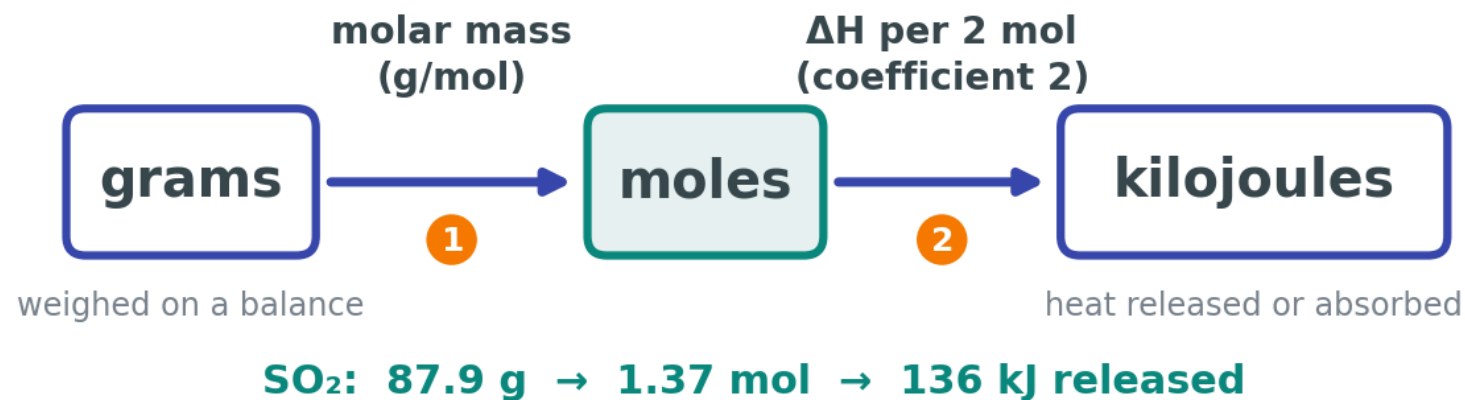
$$87.9 \text{ g} \times \frac{1 \text{ mol SO}_2}{64.07 \text{ g}} \times \frac{-198.2 \text{ kJ}}{2 \text{ mol SO}_2} = -136 \text{ kJ}$$

🧠 **Sense check:** 87.9 g is about 1.4 mol of SO_2 , each releasing about 99 kJ: near -140 kJ . ✓

✓ Worked example 4: the route on the map



given: 87.9 g SO_2 · found: -136 kJ



🧠 **Sense check:** The same two arrows as the methane route. Only the ΔH factor changes: $-198.2 \text{ kJ per } 2 \text{ mol SO}_2$. ✓

Take-home: divide by the coefficient

$$-198.2 \text{ kJ per 2 mol SO}_2 = -99.1 \text{ kJ per 1 mol SO}_2$$

the kJ-per-mole factor for any species: ΔH over that species' coefficient

ΔH belongs to the whole equation, not to any one substance. The conversion factor puts the species' coefficient under ΔH . A coefficient of 1 hides the division; a 2 does not.

Where this goes wrong

from $\text{H}_2 + \frac{1}{2} \text{O}_2 \rightarrow \text{H}_2\text{O}(\text{l}), \Delta\text{H} = -286 \text{ kJ}$, find $2 \text{H}_2\text{O}(\text{l}) \rightarrow 2 \text{H}_2 + \text{O}_2$
correct: reverse and double $\rightarrow \Delta\text{H} = +572 \text{ kJ}$

Forgetting to flip the sign. Scaling but keeping the original sign gives -572 kJ . The target is the reverse, so the sign must flip: $+572 \text{ kJ}$.

Forgetting to scale. Flipping the sign but leaving the coefficients out gives $+286 \text{ kJ}$. Every coefficient doubled, so ΔH doubles: $+572 \text{ kJ}$.

Leaving ΔH unchanged. Copying -286 kJ ignores both moves. The equation was reversed and doubled; ΔH must be too.

 Practice 3

given: this equation and ΔH · wanted: ΔH for $4 \text{NO}(\text{g}) \rightarrow 2 \text{N}_2(\text{g}) + 2 \text{O}_2(\text{g})$

What is ΔH , in kJ, for $4 \text{NO}(\text{g}) \rightarrow 2 \text{N}_2(\text{g}) + 2 \text{O}_2(\text{g})$?

- A. +360
- B. -360
- C. -180
- D. +180

✓ Practice 3 · answer: B



target: $4 \text{NO} \rightarrow 2 \text{N}_2 + 2 \text{O}_2$ · reversed and $\times 2$

$$\Delta\text{H} = -(2 \times 180 \text{ kJ}) = \mathbf{-360 \text{ kJ (answer B)}}$$

A scaled but kept the sign: $2 \times 180 = +360 \text{ kJ}$. C flipped the sign but did not scale: -180 kJ . D left ΔH untouched at $+180 \text{ kJ}$.

🧠 **Sense check:** Forming 2 mol NO absorbed 180 kJ, so making 4 mol the reverse way releases twice that: -360 kJ .





Practice 4: grams of H₂ to kilojoules



given: 1.00 g H₂ burned · wanted: ΔH in kJ

A balloon holds **1.00 g** of hydrogen gas. What is ΔH, in kJ, when it burns completely?

- A. -284
- B. +142
- C. 0.496
- D. -142

✓ Practice 4 · answer: D



$$1.00 \text{ g H}_2 \cdot 2.016 \text{ g/mol} \cdot -572 \text{ kJ per 2 mol H}_2$$

$$1.00 \text{ g} \times \frac{1 \text{ mol H}_2}{2.016 \text{ g}} \times \frac{-572 \text{ kJ}}{2 \text{ mol H}_2} = -142 \text{ kJ (answer D)}$$

A kept $-572 \text{ kJ per 1 mol H}_2$: $0.496 \times -572 = -284 \text{ kJ}$, double the answer, because the equation burns 2 mol . B dropped the sign: burning releases heat, so ΔH is negative. C stopped at moles: $1.00 \div 2.016 = 0.496 \text{ mol}$, one factor short of kilojoules.

 **Sense check:** 1.00 g is about half a mole of H_2 , and each mole releases 286 kJ : about 143 kJ released. ✓

 Practice 5

given: 2.50×10^2 kJ released · molar mass Mg 24.31 g/mol · wanted: g Mg

A magnesium flare burns white-hot. What mass of magnesium, in grams, must burn to release 2.50×10^2 kJ?

- A. 10.1
- B. 5.05
- C. 0.416
- D. 0.0171

✓ Practice 5 · answer: A



$$2.50 \times 10^2 \text{ kJ released} \cdot 1203 \text{ kJ per 2 mol Mg} \cdot \text{Mg} = 24.31 \text{ g/mol}$$

$$2.50 \times 10^2 \text{ kJ} \times \frac{2 \text{ mol Mg}}{1203 \text{ kJ}} \times \frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = \mathbf{10.1 \text{ g Mg (answer A)}}$$

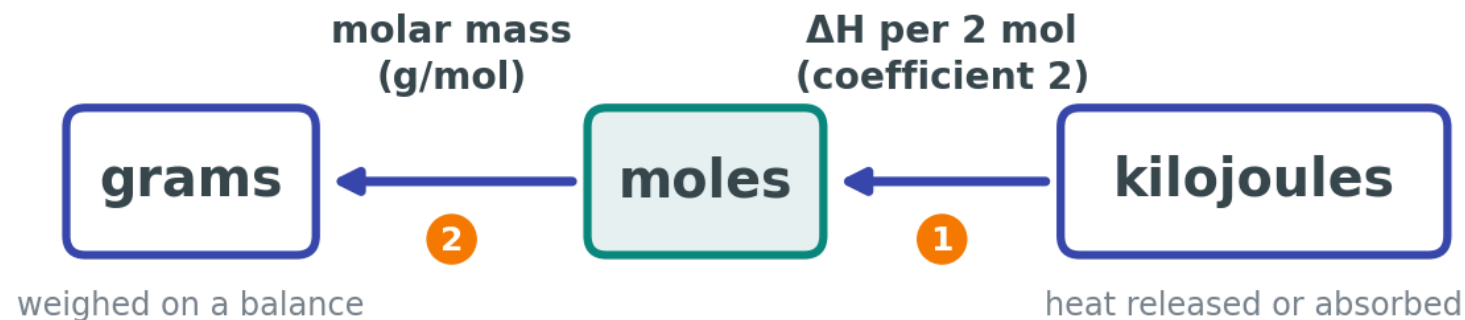
B kept 1203 kJ per 1 mol Mg: $2.50 \times 10^2 \div 1203 \times 24.31 = 5.05 \text{ g}$, half the true mass, because the equation burns 2 mol. C stopped at moles: $2.50 \times 10^2 \times 2/1203 = 0.416 \text{ mol Mg}$, one factor short of grams. D divided by the molar mass instead of multiplying: $0.416 \div 24.31 = 0.0171 \text{ g}$.

🧠 **Sense check:** One mole of Mg releases about 600 kJ, so 250 kJ takes about 0.4 mol, near 10 g at 24 g each. ✓

✓ Practice 5: the route on the map



given: $2.50 \times 10^2 \text{ kJ}$ released · found: 10.1 g Mg



Mg: 250 kJ released → 0.416 mol → 10.1 g

🧠 **Sense check:** The route runs backward. Kilojoules to moles uses ΔH written 2 mol Mg over 1203 kJ; moles to grams uses molar mass. ✓

 Practice 6

given: 35.0 g octane burned · molar mass $\text{C}_8\text{H}_{18} = 114.22 \text{ g/mol}$ · wanted: ΔH in kJ, sign included

A lawn-mower engine burns **35.0 g** of octane. What is ΔH , in kJ, for burning this amount?

- A. -5.60×10^{-5}
- B. 0.306
- C. $+1.68 \times 10^3$
- D. -1.68×10^3
- E. -3.35×10^3

✓ Practice 6 · answer: D



$$35.0 \text{ g C}_8\text{H}_{18} \cdot 114.22 \text{ g/mol} \cdot -10,941 \text{ kJ per 2 mol C}_8\text{H}_{18}$$

$$35.0 \text{ g} \times \frac{1 \text{ mol C}_8\text{H}_{18}}{114.22 \text{ g}} \times \frac{-10,941 \text{ kJ}}{2 \text{ mol C}_8\text{H}_{18}} = -1.68 \times 10^3 \text{ kJ (answer D)}$$

A flipped the ΔH factor: $0.306 \text{ mol} \times (2 \text{ mol} / -10,941 \text{ kJ}) = -5.60 \times 10^{-5}$, and mol never cancels. B stopped at moles: $35.0 \div 114.22 = 0.306 \text{ mol}$, one factor short of kilojoules. C dropped the sign: combustion releases heat, so ΔH is negative. E used $-10,941 \text{ kJ}$ per 1 mol: $0.306 \times (-10,941) = -3.35 \times 10^3 \text{ kJ}$, double the answer, because the equation burns 2 mol.

 **Sense check:** Per mole of octane the heat is $-10,941 \div 2 = -5470.5 \text{ kJ}$, and 0.306 mol is under a third of a mole: about 1680 kJ released. ✓

Check yourself

1. An equation is reversed and its coefficients tripled. What two changes apply to ΔH ?
2. A reaction releases 500 kJ per mole. Write the conversion factor that turns moles of it into kilojoules, and state which unit cancels.

The kJ-per-mole factor returns with melting and boiling. A heat of fusion or vaporization in kJ/mol turns moles of a substance into kilojoules, along the same grams $\rightarrow$ moles $\rightarrow$ kilojoules route.

Can you...?

- ☐ tell kinetic from potential energy and convert energy among joules, calories, kilojoules, and food Calories?
- ☐ classify a process as exothermic or endothermic from the sign of ΔH , a heat term in the equation, or what is observed?
- ☐ use $q = m \cdot c \cdot \Delta T$ to find heat, mass, specific heat, or temperature change, and compare how specific heats set the temperature response to the same heat?
- ☐ apply heat lost = heat gained in a coffee-cup calorimeter to find the heat change of a metal, its specific heat, or the heat of a reaction in solution?
- ☐ find the energy content of a food from the water's temperature rise in a bomb calorimeter, in kJ and Calories per gram?
- ☐ write a thermochemical equation and use ΔH as a conversion factor for moles or grams of any substance in it?

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