

Thermochemistry

CHEM 50 · Dr. Karmach

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

- Distinguish heat from work and convert among joules, calories, and Calories
- Identify exothermic and endothermic processes from equations, energy diagrams, and the sign of ΔH
- Apply $q = m \cdot c \cdot \Delta T$ to find heat, mass, specific heat, or temperature change
- Compare how specific heats set different substances' temperature response to the same heat
- Use calorimetry data to determine the heat of a process or reaction
- Reverse and scale thermochemical equations and use ΔH as a conversion factor
- Combine reversed and scaled equations with Hess's law to find an overall reaction's enthalpy
- Write formation equations and use standard formation enthalpies to find any reaction's enthalpy
- Compute heat for melting, vaporizing, or a full heating curve

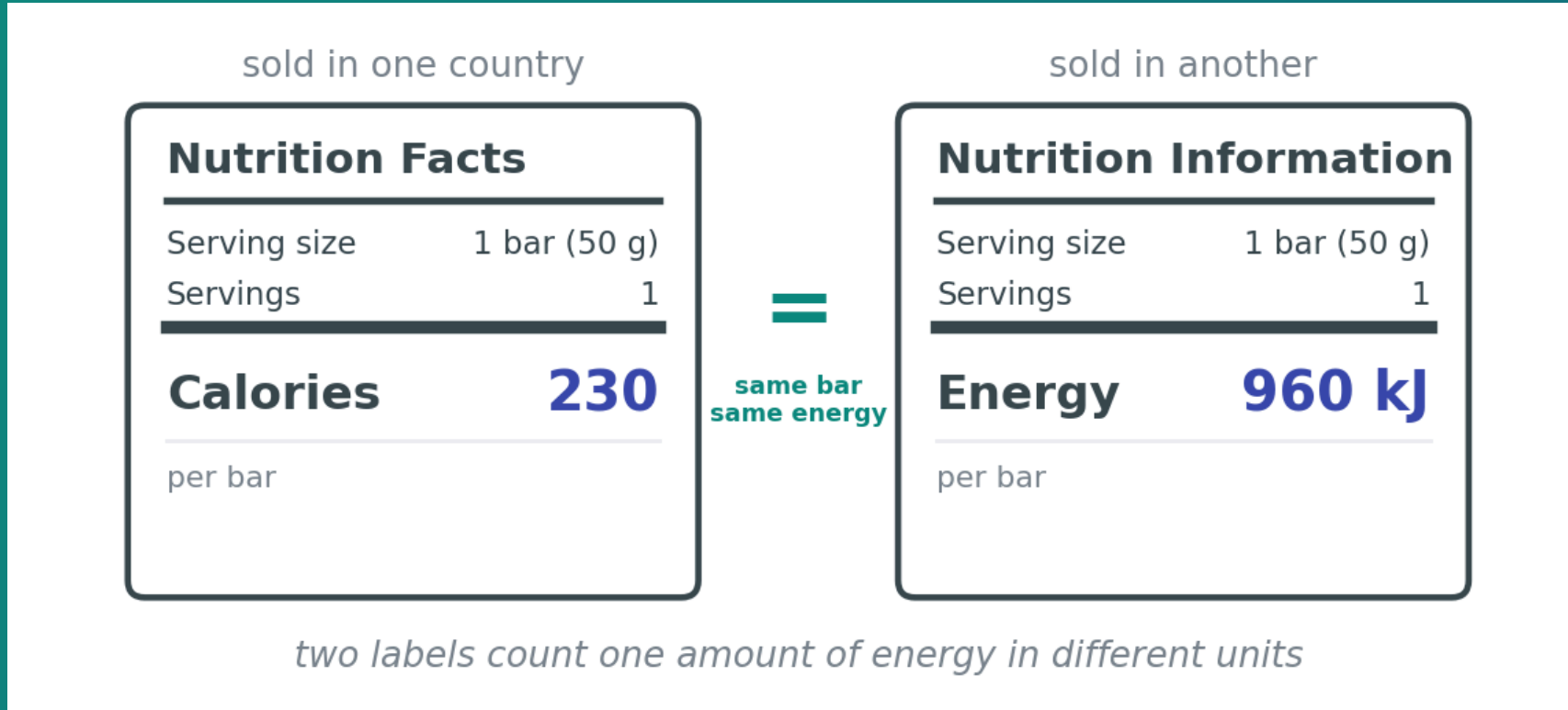
Today's route

1. **Energy & Its Units**
2. **Internal Energy & the First Law**
3. **Exothermic & Endothermic**
4. **Specific Heat & $q = mc\Delta T$**
5. **Comparing Specific Heats**
6. **Heat of a Phase Change**
7. **Heating & Cooling Curves**
8. **Calorimetry**
9. **Thermochemical Equations**
10. **Hess's Law**
11. **Standard Enthalpies of Formation**

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



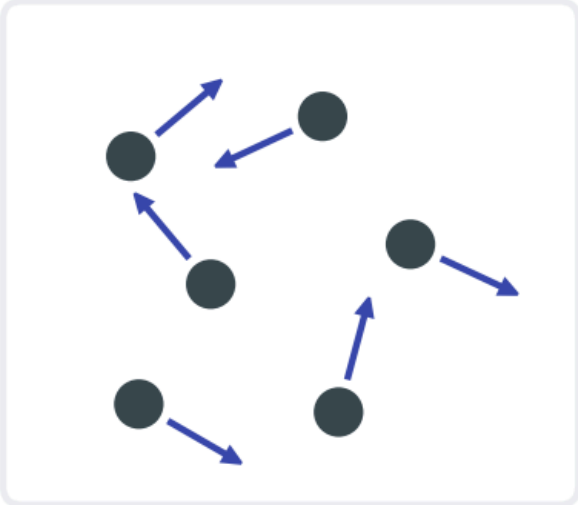
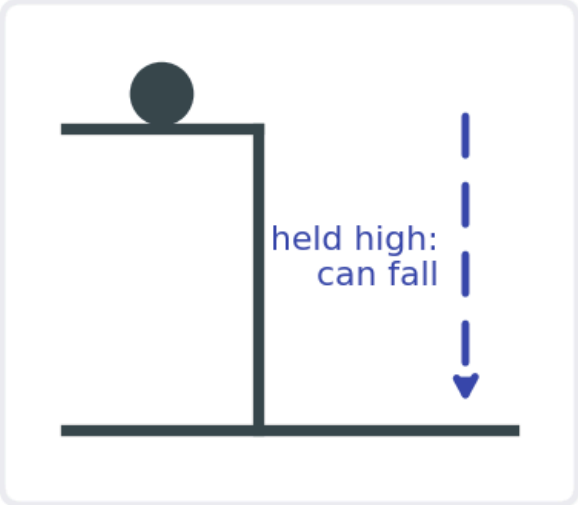
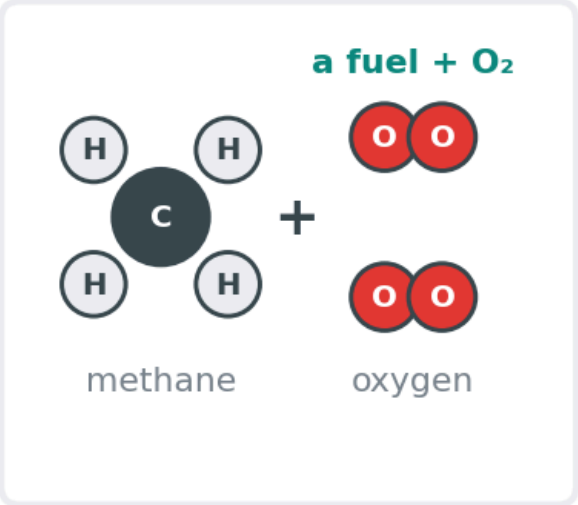
The same snack bar sells in two countries. One label lists **230 Calories**; the other lists **960 kJ**. Both describe the same energy, counted in different units.

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

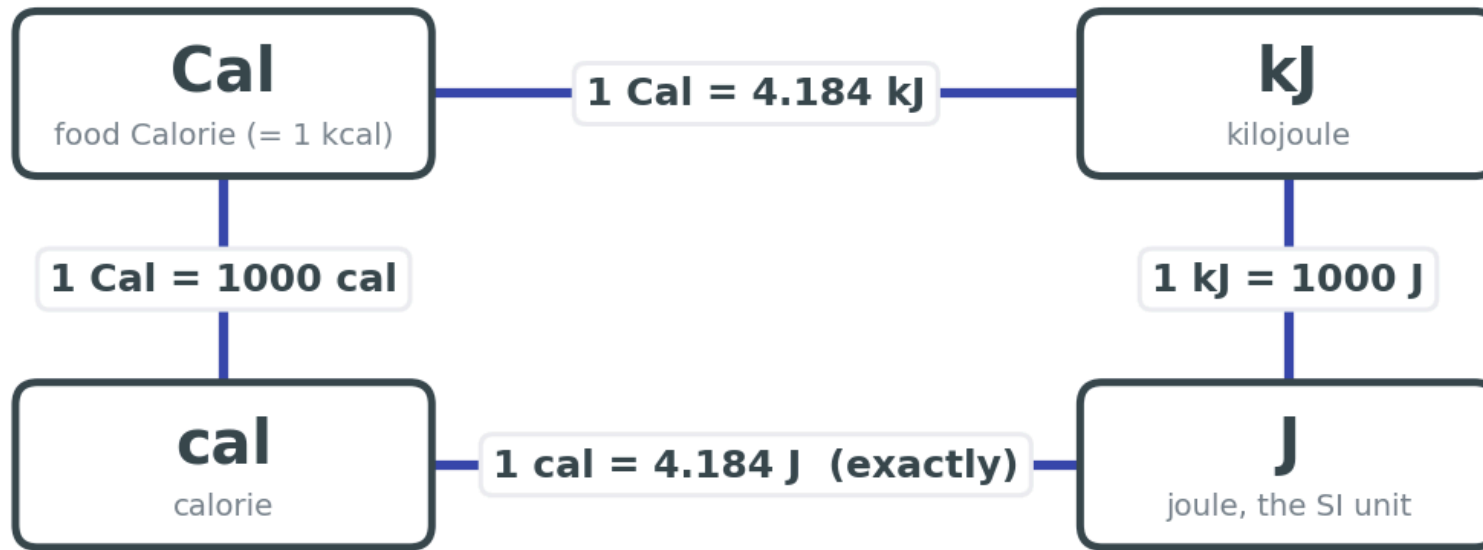
work (w): expanding gas pushes a piston
energy moves because a force acts through a distance

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

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

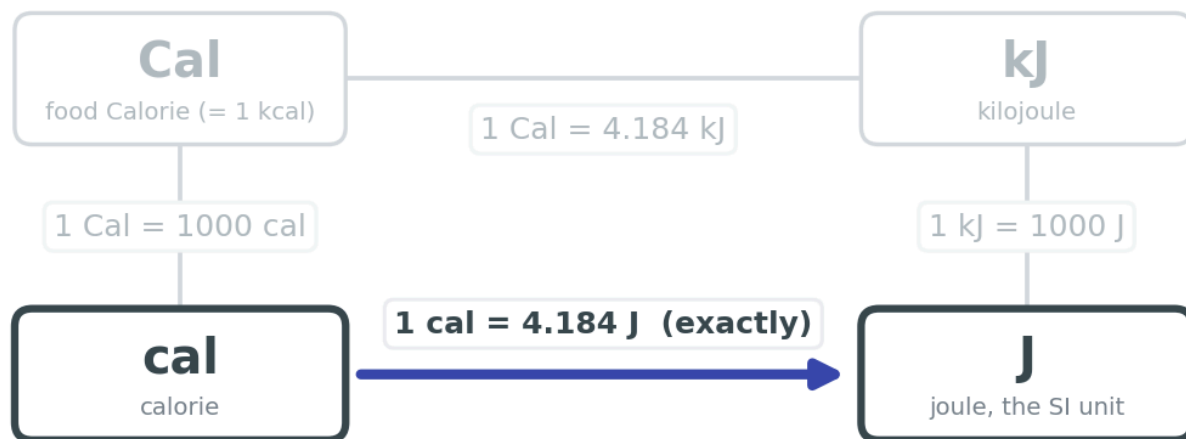
Step 4 · Sense-check

 **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

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

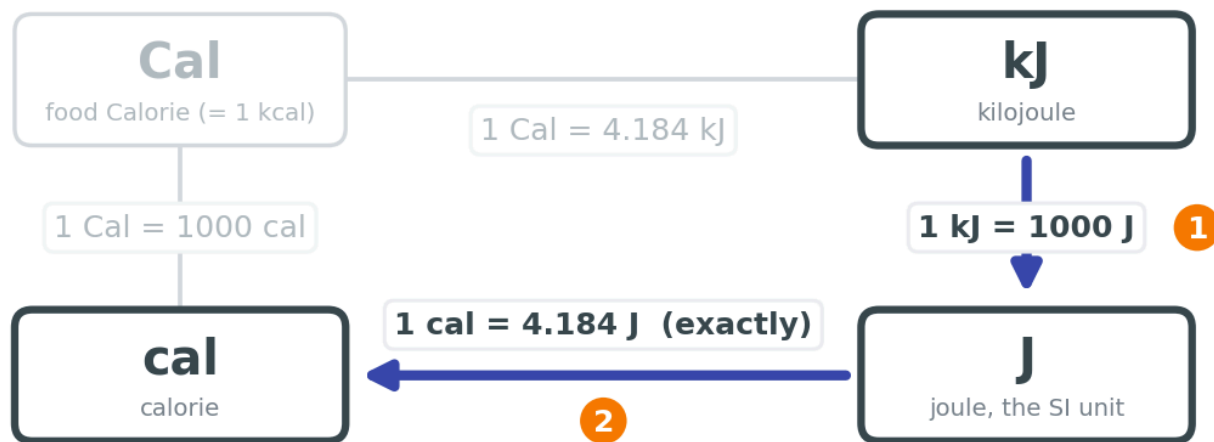
Step 4 · Sense-check

 **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

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

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 1 · 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

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

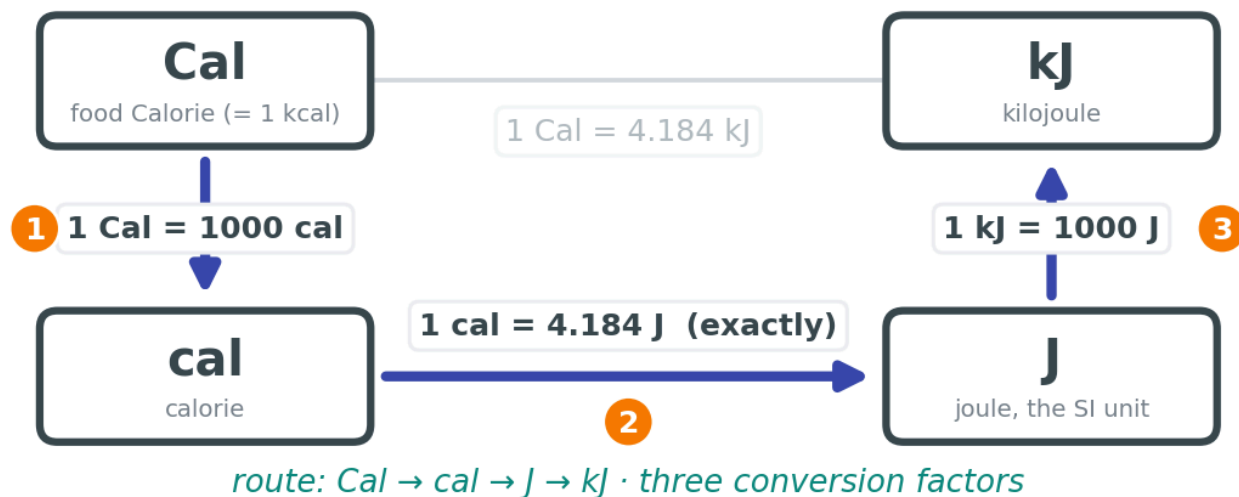
Step 4 · Sense-check

 **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 2

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

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 2 · 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

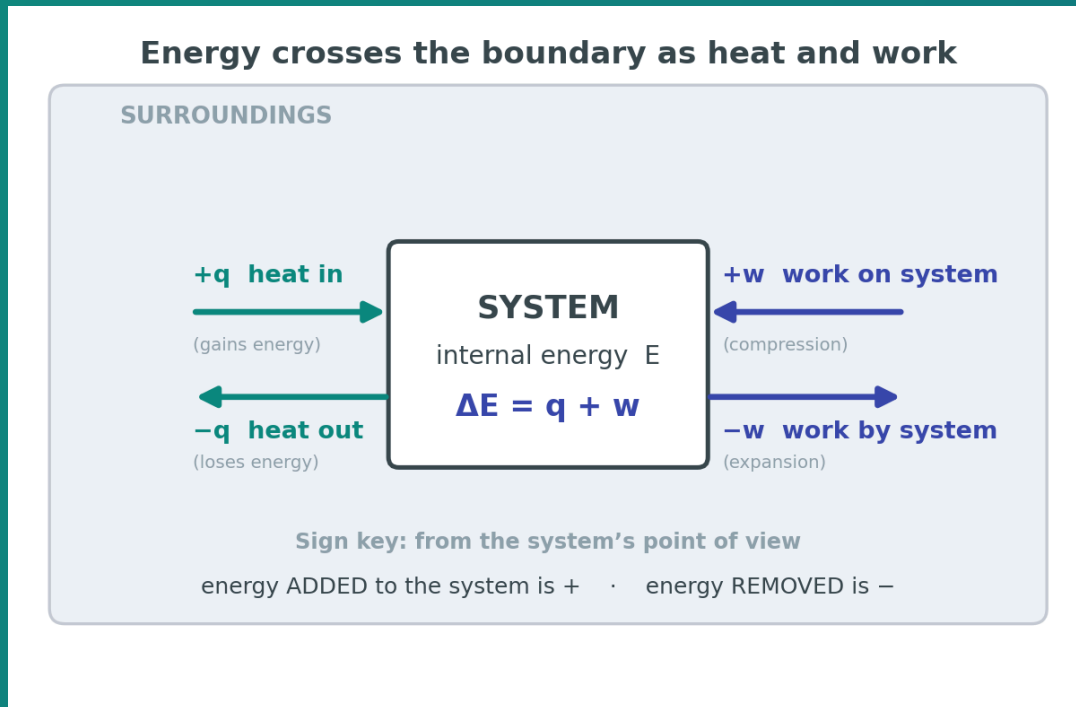
1. A label lists **95 Calories**. Write the chain to kilojoules: which factor comes first, and what does each unit cancel into?
2. A cup of water and a pot of water both read 60 °C. Which quantity matches, and which differs: temperature, or energy content?

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 · Internal Energy & the First Law

Track a system's internal energy with the first law, $\Delta E = q + w$, assigning the correct signs to heat and work, computing pressure-volume work from $w = -P \cdot \Delta V$, and telling a state function (E , ΔE) from a path function (q , w).

Energy has nowhere to hide



A system can gain or lose energy, but the total never vanishes. Every joule it loses turns up in the surroundings, crossing the boundary as heat or work.

Internal energy: the system's whole energy account

The **internal energy E** is the total of every kinetic and potential energy inside the system: all the molecular motion and every bond and interaction, added together.

E = all kinetic energy + all potential energy inside the system

its absolute size is out of reach: thermochemistry tracks only the change, $\Delta E = E(\text{final}) - E(\text{initial})$

The first law: energy is conserved

Energy is never created or destroyed: it only moves. A system's gains come from the surroundings; its losses go to them. It crosses the boundary as **heat (q)** or **work (w)**.

$$\Delta E = q + w$$

the first law of thermodynamics: a system's energy change equals the heat plus the work exchanged with its surroundings

Open, closed, and isolated systems

open system: matter and energy both cross · an open pot of soup
steam escapes (matter) · heat escapes (energy)

closed system: energy crosses, matter cannot · a sealed flask
the glass passes heat · nothing inside gets out

isolated system: neither crosses · a stoppered vacuum thermos, nearly
no matter out · almost no heat out

Count what can cross the boundary: an open system passes 2 things, a closed system 1, an isolated system 0.

Adding signed amounts

$$(+200) + (-50) = +150$$

a deposit of 200 and a withdrawal of 50 · money in is +, money out is – · the balance rises by 150

A bank statement signs every entry from the account's side, then adds. Energy is tracked the same way: heat and work are the entries, and ΔE is the change in the balance.

One sign rule: positive means energy in

Give q and w a **+ sign when they add energy to the system**, a **– sign when they take energy out**. Read every arrow from the system's point of view.

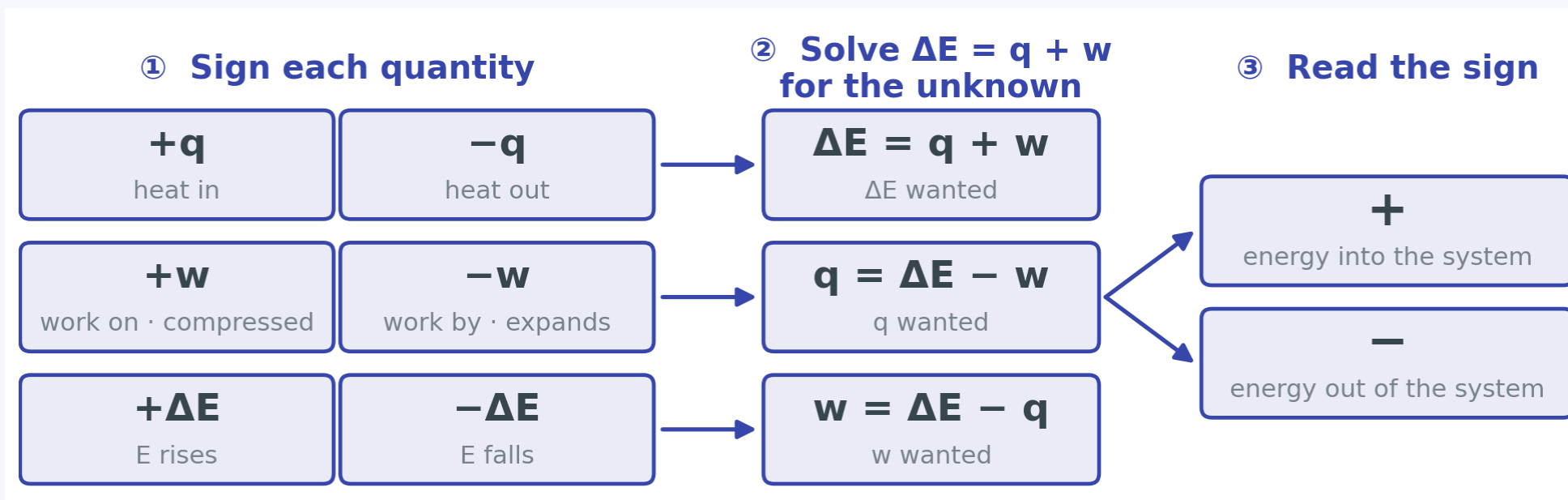
+ q heat INTO system · – q heat OUT · + w work done ON system · – w work done BY system

endothermic: $q > 0$ · exothermic: $q < 0$ · compression: $w > 0$ · expansion: $w < 0$



The method

1. **Sign each quantity:** heat in $+q$, out $-q$; work on $+w$, by $-w$.
2. **Solve $\Delta E = q + w$ for the unknown:** ΔE , q , or w .
3. **Read the sign:** positive means energy in.





Guided example: a refrigerator compressor

310 J of work done on the gas · 450 J of heat released by the gas
system: the coolant gas · wanted: ΔE of the gas

In a refrigerator, the compressor squeezes the coolant gas with **310 J of work**. The hot gas then releases **450 J of heat** to the kitchen air.

Find ΔE for the gas.

✓ Guided example: solution

Step 1 · Sign each quantity

Step 2 · Solve $\Delta E = q + w$ for the unknown

$$\Delta E = q + w = (-450 \text{ J}) + (+310 \text{ J}) = -140 \text{ J}$$

q : heat out, $-$ · w : work on the gas, $+$ · given: 310 J of work on, 450 J of heat out

Step 3 · Read the sign

① Sign each quantity

② Solve for the unknown

③ Read the sign



 **Sense check:** The gas lost 140 J: the 450 J of heat out outweighs the 310 J of work in. ✓



Practice 1

$$\Delta E = q + w$$

given: 165 J of heat lost · 58 J of work done by the gas · wanted: ΔE of the gas

Hot gas in an engine cylinder pushes its piston outward, doing 58 J of work, while it loses 165 J of heat to the cylinder walls. What is ΔE for the gas, in joules?

- A. +107
- B. -223
- C. +223
- D. -107

✓ Practice 1 · answer: B

$$\Delta E = q + w = (-165 \text{ J}) + (-58 \text{ J}) = -223 \text{ J}$$

heat lost: $q = -165 \text{ J}$ · work done by the gas, an expansion: $w = -58 \text{ J}$ · answer B

A flipped the heat sign: $(+165) + (-58) = +107 \text{ J}$. C reversed both signs: $165 + 58 = +223 \text{ J}$, the surroundings' ΔE , not the gas's. D subtracted the work: $(-165) - (-58) = -107 \text{ J}$.

🧠 **Sense check:** Both flows carry energy out of the gas, so both terms are negative and so is their sum. ✓

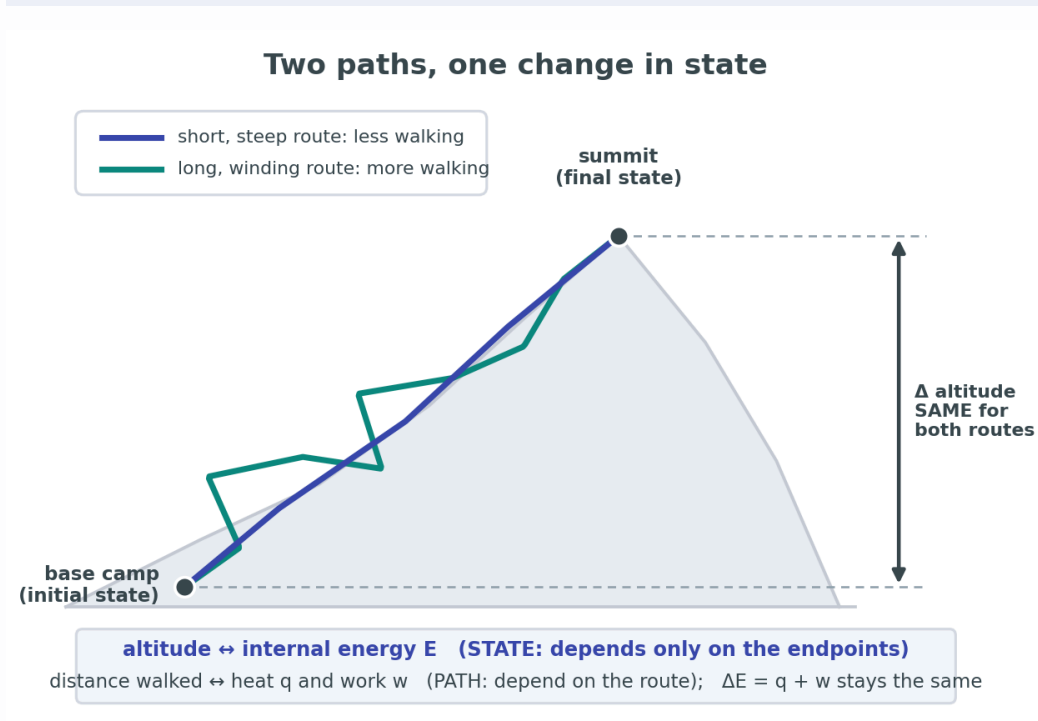
① Sign each quantity

② Solve for the unknown

③ Read the sign



State functions vs path functions



E and ΔE are state functions: they depend only on the initial and final states, never on the route. **q and w are path functions:** their values depend on the route.

altitude ↔ E (state) · distance walked ↔ q and w (path)
two routes share one ΔE , yet split it into different amounts of q and w



Worked example: finding the heat

$$\Delta E = q + w$$

given: $\Delta E = -150 \text{ J}$ · 240 J of work done ON the system · wanted: q

A system's internal energy drops by 150 J even though the surroundings do 240 J of work on it. Find the heat, with its sign.

✓ Worked example: solution

$$\Delta E = q + w$$

given: $\Delta E = -150 \text{ J}$ · 240 J of work done ON the system · wanted: q

Step 1 · Sign each quantity

Step 2 · Solve $\Delta E = q + w$ for the unknown

$$q = \Delta E - w = (-150 \text{ J}) - (+240 \text{ J}) = -390 \text{ J}$$

the system released 390 J of heat

Step 3 · Read the sign

 **Sense check:** The system gained 240 J by work yet still ended 150 J down, so heat had to carry away both amounts: 390 J out. Negative q , heat released. ✓

① Sign each quantity

② Solve for the unknown

③ Read the sign

$$+w$$

$$w = +240 \text{ J}$$

$$-\Delta E$$

$$\Delta E = -150 \text{ J}$$

$$q = \Delta E - w$$

$$q = -390 \text{ J}$$

$$-$$

$$390 \text{ J of heat out}$$

Your turn: finding the work

$$\Delta E = q + w$$

given: $q = +425 \text{ J}$ absorbed · $\Delta E = +125 \text{ J}$ · wanted: w

$$w = \Delta E - q = (\text{ } \text{J}) - (\text{ } \text{J}) = \text{ } \text{J}$$

sign the two givens, subtract, and read the result's sign

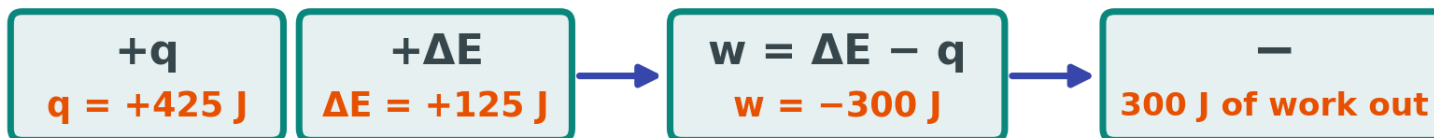
$$w = \Delta E - q = (+125 \text{ J}) - (+425 \text{ J}) = -300 \text{ J}$$

w negative: the gas did 300 J of work on the surroundings · an expansion

① Sign each quantity

② Solve for the unknown

③ Read the sign



 Practice 2

$$\Delta E = q + w$$

given: 75 J of heat given off · internal energy up by 115 J · wanted: w

Steam in a sealed cylinder gives off 75 J of heat while its internal energy goes up by 115 J. What is w for the steam, in joules?

- A. -190
- B. +40
- C. +190
- D. +75

✓ Practice 2 · answer: C

$$w = \Delta E - q = (+115 \text{ J}) - (-75 \text{ J}) = +190 \text{ J}$$

heat given off: $q = -75 \text{ J}$ · energy went up: $\Delta E = +115 \text{ J}$ · answer C

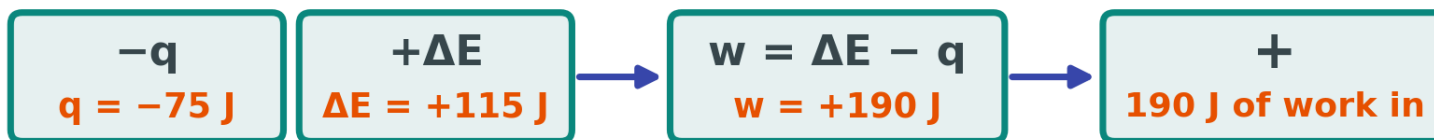
A subtracted in reverse: $q - \Delta E = (-75) - (+115) = -190 \text{ J}$, the work read from the surroundings' side. B flipped the heat sign: $115 - 75 = +40 \text{ J}$, as if the steam had absorbed the heat. D treated ΔE as zero: $w = -q = +75 \text{ J}$, true only for a cycle back to the start.

 **Sense check:** The steam lost 75 J as heat yet ended 115 J up, so work brought in both: $75 + 115 = 190 \text{ J}$. Positive w : the steam was compressed. ✓

① Sign each quantity

② Solve for the unknown

③ Read the sign





Practice 3

$$\Delta E = q + w$$

given: path 1: $q = +170 \text{ J}$, $w = +50 \text{ J}$ · path 2: $q = +310 \text{ J}$ · wanted: w on path 2

A gas is taken between the same two states along two paths. On path 1 it absorbs 170 J of heat while the surroundings do 50 J of work on it. On path 2 it absorbs 310 J of heat.

What is w on path 2, in J?

- A. +530
- B. +220
- C. +50
- D. -90

✓ Practice 3 · answer: D

$$\text{path 1: } \Delta E = q + w = (+170 \text{ J}) + (+50 \text{ J}) = +220 \text{ J}$$

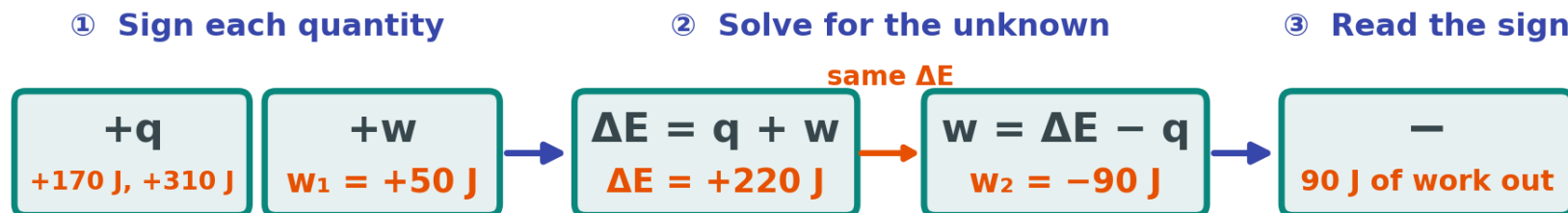
E is a state function: path 2 starts and ends in the same states, so its ΔE is also +220 J

$$\text{path 2: } w = \Delta E - q = (+220 \text{ J}) - (+310 \text{ J}) = -90 \text{ J}$$

answer D: w negative, so on path 2 the gas does 90 J of work on the surroundings

A flipped the sign of path 2's heat: $220 + 310 = +530 \text{ J}$, as if the 310 J had left the gas. B stopped at ΔE : +220 J is the energy change both paths share, not path 2's work. C copied path 1's work, +50 J: w is a path function and changes with the route.

🧠 **Sense check:** Path 2 brings in 310 J of heat but may keep only 220 J, so 90 J must leave as work. ✓



ΔH : the heat at constant pressure

$$q_p = \Delta H$$

constant pressure: an open flask on the bench · the heat measured there is the enthalpy change

Run a process open to the air and its heat is the **enthalpy change, ΔH** . Most chemistry runs at constant pressure, which is why ΔH is the number tables report.

⚠ Where this goes wrong

Adding magnitudes, ignoring signs. Work done BY the system is $-w$. A gas that absorbs 125 J and does 78 J of work has $\Delta E = 125 - 78 = 47$ J, not 203 J.

Flipping the heat sign. Heat absorbed is $+q$ (endothermic); heat released is $-q$ (exothermic). Reading "releases" as $+$ gives the wrong ΔE , often the wrong sign entirely.

Signing the work by the wrong party. "The surroundings do 240 J of work ON the system" means $w = +240$ J. Read it as work done BY the system and w flips to -240 J, landing ΔE a full 480 J off. Every sign reads from the system's side.

Calling q or w a state function. Only E and ΔE are path-independent. Heat and work depend on the route: one ΔE can come from many different q, w pairs.



Extra example 1: heat in, work out

absorbs 125 J of heat · does 78 J of work on the surroundings

wanted: ΔE of the system

A gas in a cylinder **absorbs 125 J of heat**, then **does 78 J of work** pushing its piston outward. Find the change in its internal energy.

✓ Extra example 1: solution

Step 1 • Sign each quantity

Step 2 • Solve $\Delta E = q + w$ for the unknown

$$\Delta E = q + w = (+125 \text{ J}) + (-78 \text{ J}) = +47 \text{ J}$$

energy in (125) outweighs energy out (78), so E rises by 47 J

Step 3 • Read the sign

🧠 **Sense check:** Watch the signs: adding the magnitudes ($125 + 78 = 203 \text{ J}$) would treat the expansion as energy gained. The system spent 78 J doing work, so it keeps only 47 J of the 125 J it took in. ✓

① Sign each quantity

② Solve for the unknown

③ Read the sign



Work from a changing volume

$$w = -P \cdot \Delta V$$

P: the constant external pressure, in atm · $\Delta V = V(\text{final}) - V(\text{initial})$, in L · 1 L·atm = 101.325 J

A gas that **expands** pushes the surroundings back: it spends energy, w negative. A gas that is **compressed** receives that push: w positive. The minus sign in the formula produces both results automatically.



Extra example 2: work of an expanding gas

$$w = -P \cdot \Delta V \cdot 1 \text{ L} \cdot \text{atm} = 101.325 \text{ J}$$

given: a gas expands from 264 mL to 971 mL · wanted: w (a) against a vacuum (b) against a constant 4.0 atm

A gas expands from **264 mL to 971 mL** at constant temperature. Find the work done by the gas if it expands (a) against a vacuum and (b) against a constant pressure of 4.0 atm.

✓ Extra example 2: solution

a gas expands from 264 mL to 971 mL

$$w = -P \cdot \Delta V \cdot 1 \text{ L} \cdot \text{atm} = 101.325 \text{ J} \cdot (\text{a) vacuum (b) 4.0 atm}$$

Part (a) • expansion against a vacuum

$$w = -P \cdot \Delta V = -(0)(0.707 \text{ L}) = 0$$

there is no pressure in a vacuum, so $w = 0$ • nothing pushes back, no work

Part (b) • expansion against 4.0 atm

$$w = -4.0 \text{ atm} \times 0.707 \text{ L} = -2.83 \text{ L} \cdot \text{atm} \times \frac{101.325 \text{ J}}{1 \text{ L} \cdot \text{atm}} = -2.9 \times 10^2 \text{ J}$$

 **Sense check:** Expansion is work done BY the gas, so w is negative: -286.5 J , kept to two significant figures as $-2.9 \times 10^2 \text{ J}$. Against a vacuum the same expansion costs nothing; work needs something to push against. ✓



Extra practice

$$w = -P \cdot \Delta V \cdot 1 \text{ L} \cdot \text{atm} = 101.325 \text{ J}$$

given: a gas is compressed from 982 mL to 351 mL by a constant 3.0 atm · wanted: w

A piston compresses a gas from **982 mL to 351 mL** under a constant external pressure of 3.0 atm. What is w for the gas, in J?

- A. $+1.9 \times 10^2$
- B. -1.9×10^2
- C. +1.9
- D. $+1.9 \times 10^{-2}$

✓ Extra practice · answer: A

$$\Delta V = 0.351 \text{ L} - 0.982 \text{ L} = -0.631 \text{ L}$$

compression: the volume falls, so ΔV is negative

$$w = -3.0 \text{ atm} \times (-0.631 \text{ L}) = +1.89 \text{ L} \cdot \text{atm} \times \frac{101.325 \text{ J}}{1 \text{ L} \cdot \text{atm}} = +1.9 \times 10^2 \text{ J (answer A)}$$

B dropped the minus sign in $w = -P \cdot \Delta V$: $3.0 \times (-0.631) \times 101.325 = -1.9 \times 10^2 \text{ J}$, which would say the squeezed gas lost energy. C stopped at $+1.89 \text{ L} \cdot \text{atm}$ and wrote it as joules; only the $101.325 \text{ J per L} \cdot \text{atm}$ factor makes joules. D turned the conversion factor upside down: $1.89 / 101.325 = 1.9 \times 10^{-2} \text{ J}$.

 **Sense check:** Compression is work done ON the gas: energy flows in, so w must be positive. The formula's minus sign and the negative ΔV cancel to say exactly that. ✓



Check yourself

1. A system releases 90 J of heat and has 140 J of work done on it. Find ΔE , and say whether its energy rose or fell.
2. A gas is taken between the same two states by two different routes. Which of q , w , and ΔE must be the same on both routes, and why?

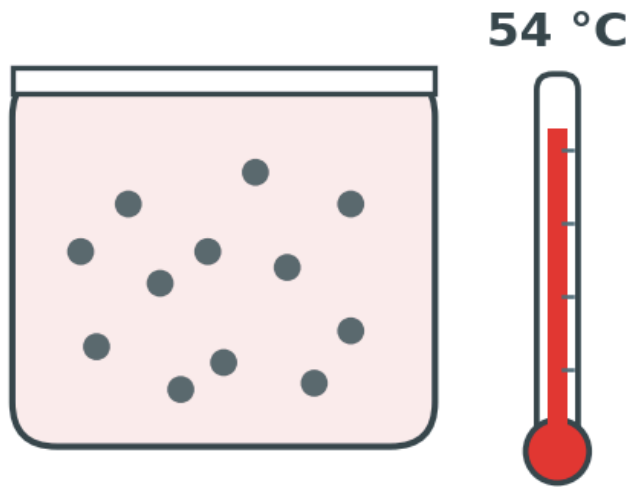
At constant pressure, the heat q gets its own name: the **enthalpy change ΔH** . Its sign sorts every process into exothermic or endothermic.

3 · 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.

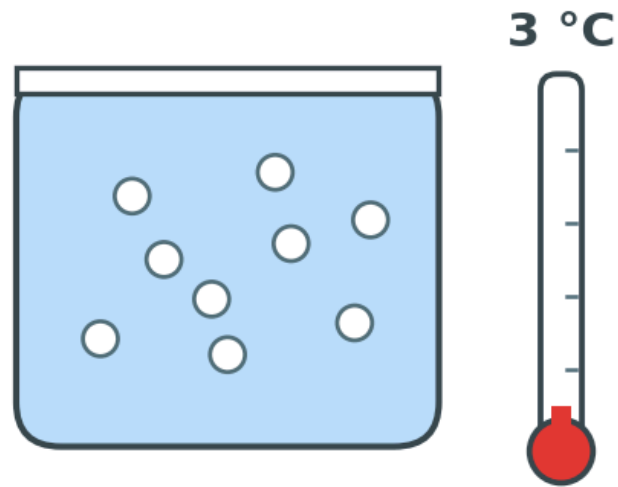
Two pouches from the drugstore

snap this pouch: it heats up



sealed pouch: iron powder + air

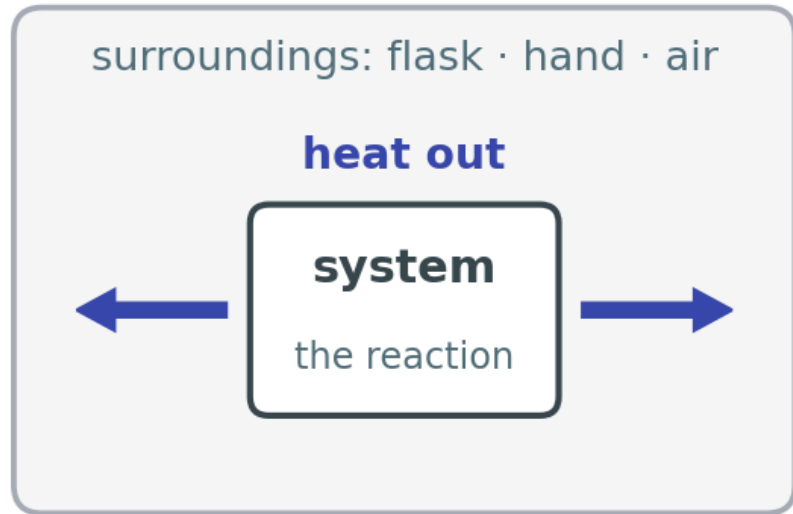
snap this pouch: it turns icy



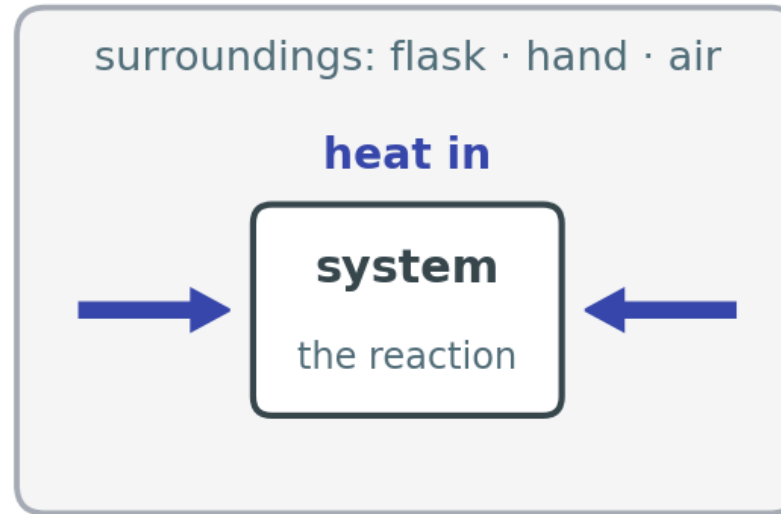
sealed pouch: water + salt pellets

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

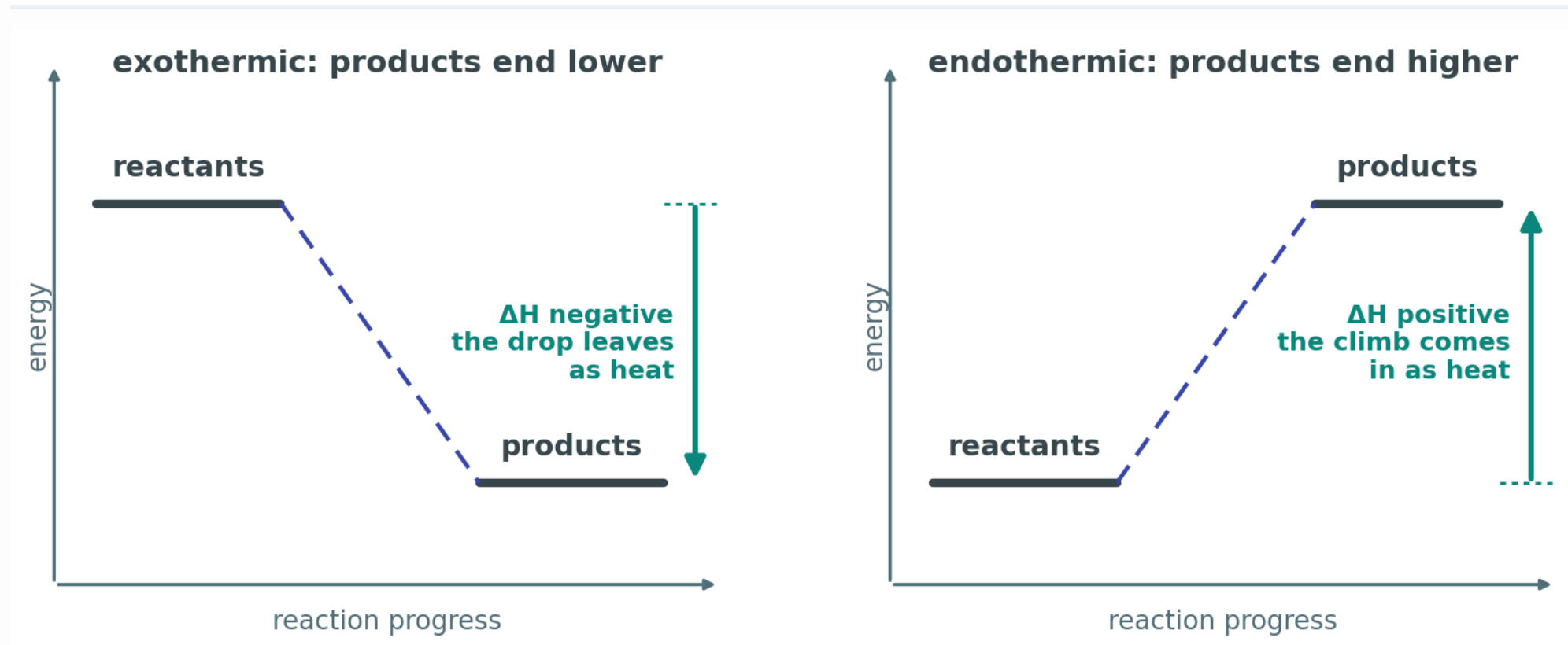
surroundings cool down · ΔH positive · melting ice, the cold-pack pouch

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

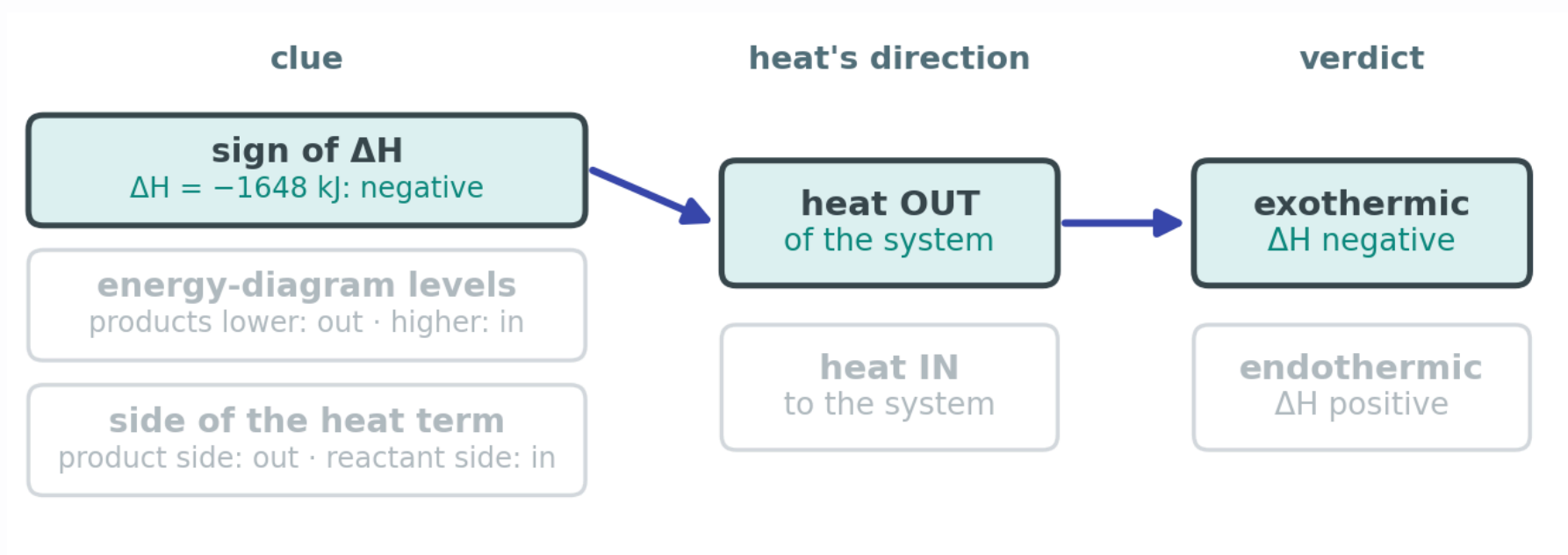
$4 \text{ Fe(s)} + 3 \text{ O}_2\text{(g)} \rightarrow 2 \text{ Fe}_2\text{O}_3\text{(s)} + 1648 \text{ kJ} \rightarrow \text{exothermic}$
 heat out · $\Delta H = -1648 \text{ kJ}$ · the surroundings (your hands) warm up

 **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

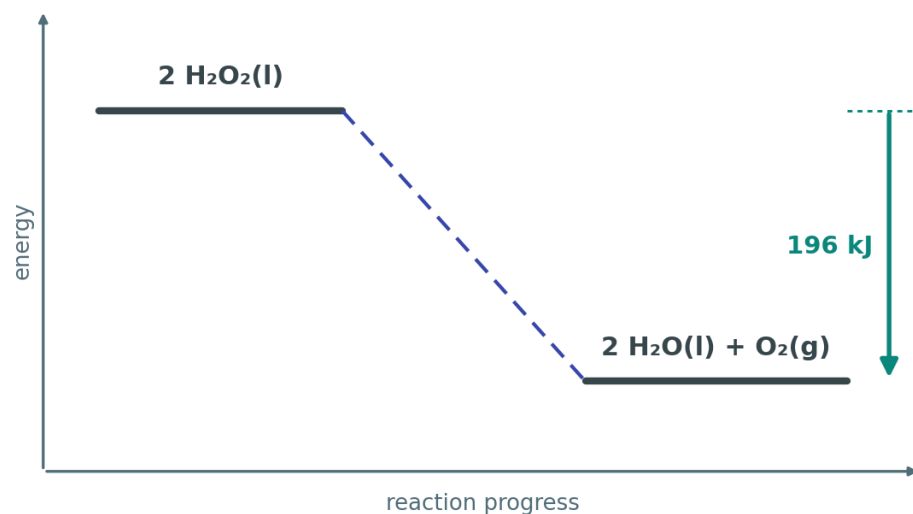


🧠 **Sense check:** Any one of the three clues gives the heat's direction. Here the sign settled it: negative, heat out, exothermic. ✓

Worked example 2: hydrogen peroxide decomposes

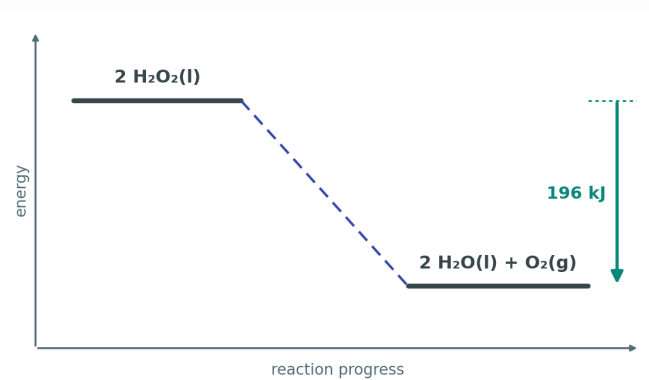


given: the energy diagram · wanted: the verdict and ΔH with its sign



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

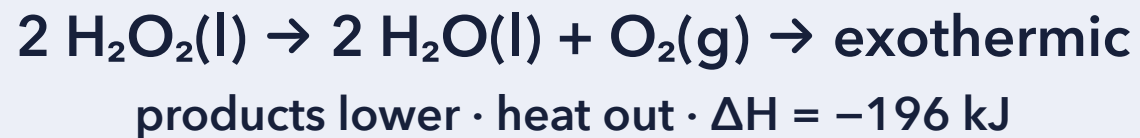
✓ Worked example 2: solution



Step 1 • Name the system

Step 2 • Find the heat's direction

Step 3 • State the verdict

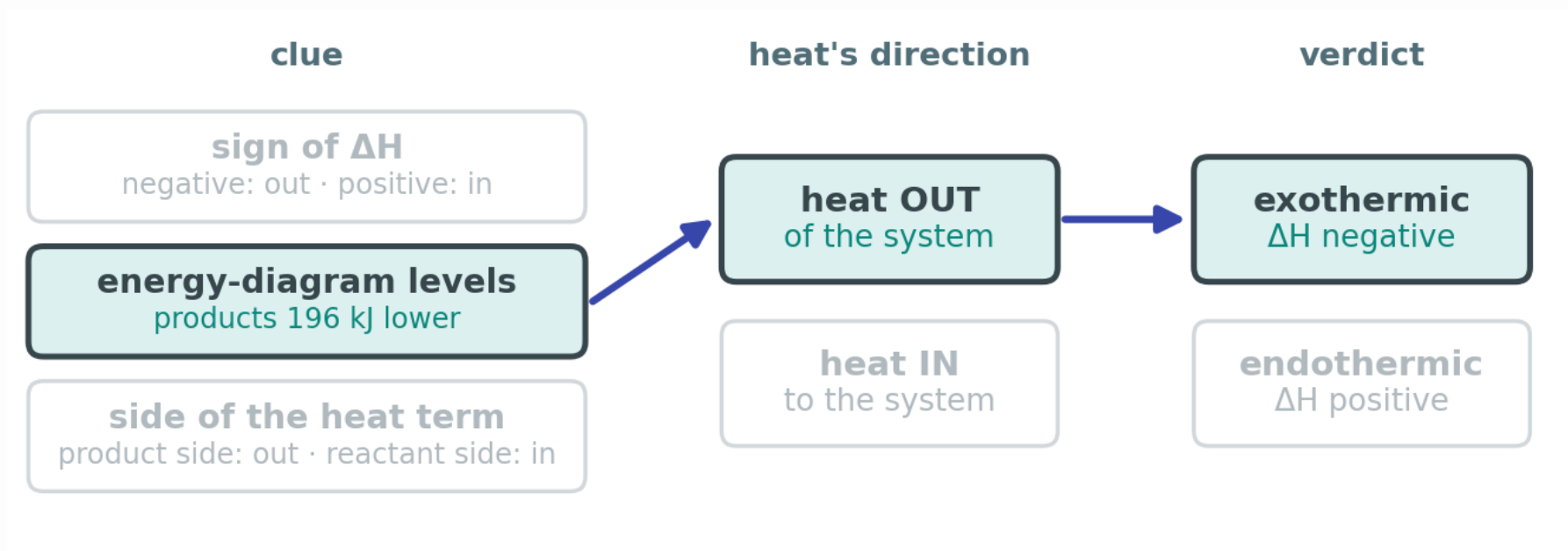


 **Sense check:** Downhill on an energy diagram is heat out. The products hold less energy than the reactants, and the fizzing cut warms slightly.

✓ 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 → endothermic

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

⚠ 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



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 2 · 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

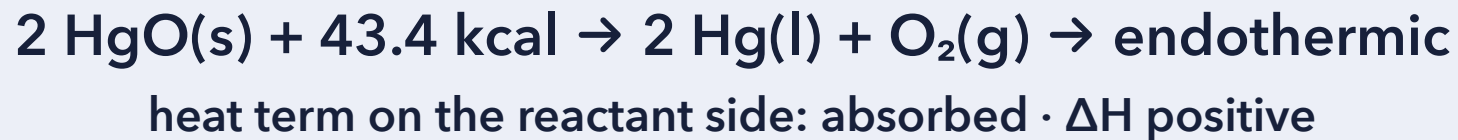


given: the heat term in kcal $\cdot$ 1 cal = 4.184 J

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: $43.4 \times 4.184 = 182$. ✓



Extra practice 3



given: 154 kJ of heat released · 2.48 kJ of work done by the H₂ pushing back the air

One mole of zinc dissolves in hydrochloric acid in an open flask. Which statement describes the reaction?

- A. Endothermic · $\Delta H = -154 \text{ kJ}$ · $\Delta E = -156 \text{ kJ}$
- B. Exothermic · $\Delta H = -154 \text{ kJ}$ · $\Delta E = -152 \text{ kJ}$
- C. Endothermic · $\Delta H = +154 \text{ kJ}$ · $\Delta E = +152 \text{ kJ}$
- D. Exothermic · $\Delta H = -154 \text{ kJ}$ · $\Delta E = -156 \text{ kJ}$

✓ Extra practice 3 · answer: D



open flask: $\Delta H = q = -154 \text{ kJ}$ · the H_2 pushes back the air, work done BY the system: $w = -2.48 \text{ kJ}$

$$\Delta E = q + w = (-154 \text{ kJ}) + (-2.48 \text{ kJ}) = -156 \text{ kJ (answer D)}$$

ΔH counts the heat alone · ΔE counts the heat and the work

A paired the label with the wrong sign: heat out is exothermic. B signed the work as done on the system: $-154 + 2.48 = -152 \text{ kJ}$. The H_2 pushed the air back, so that energy left. C read the heat sign from the surroundings: $+154 - 2.48 = +152 \text{ kJ}$. The room gains the heat; the system loses it.

 **Sense check:** Heat and work both leave the system, so ΔE is a little more negative than ΔH . Both negative: the flask warms its surroundings. ✓

Check yourself

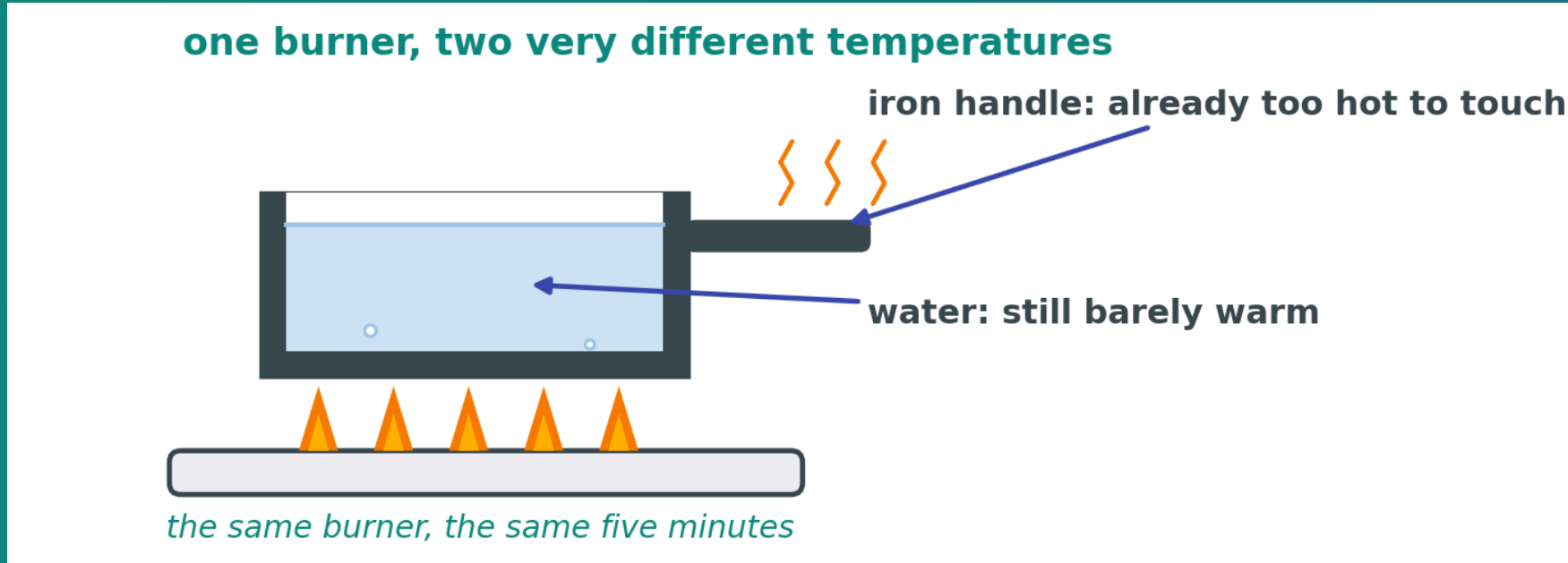
1. Propane burns in a camp stove: $\text{C}_3\text{H}_8(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 3 \text{CO}_2(\text{g}) + 4 \text{H}_2\text{O}(\text{l})$, $\Delta H = -2220$ kJ. State the verdict and which way heat flows.
2. An energy diagram shows the products 178 kJ above the reactants. Give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

4 · 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

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

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 · $c = 0.897 \text{ J/g} \cdot ^\circ\text{C}$ · $T_{\text{initial}} = 95.0 \text{ }^\circ\text{C}$ · 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}

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

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

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

↓ 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: 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 3 · 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 4

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

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 4 · answer: C

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

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

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

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: $157.0 \text{ } ^\circ\text{C}$.

 **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

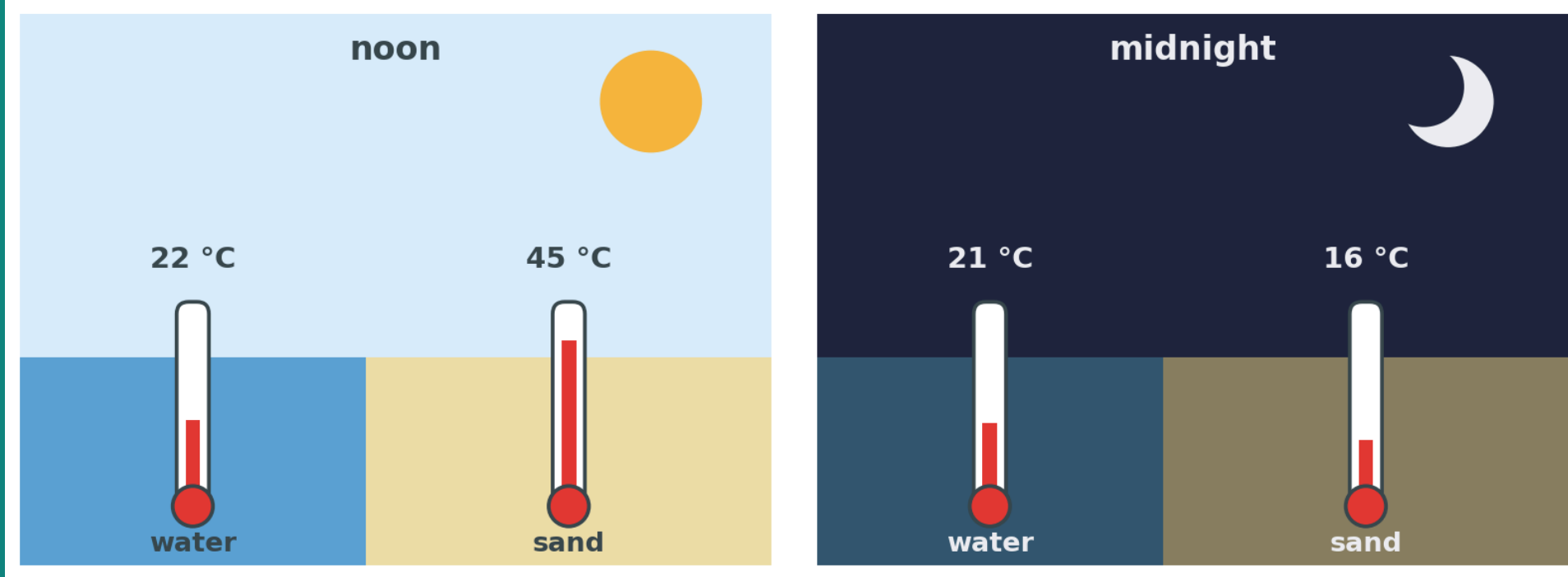
1. Water cools from 50.0 °C to 20.0 °C. Write ΔT with its sign. What is the sign of q , and what does it say about the heat?
2. Solve $q = m \cdot c \cdot \Delta T$ for m , in symbols. Which units cancel, and which unit survives?

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 .

5 · 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.

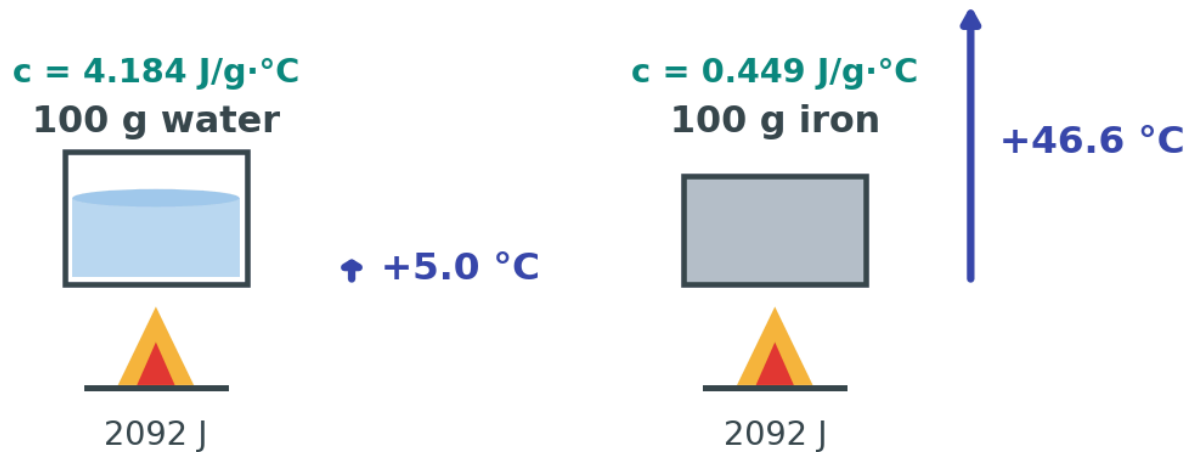
The beach, twelve hours apart



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.

Same heat and mass: c and ΔT trade off

the same heat into the same mass

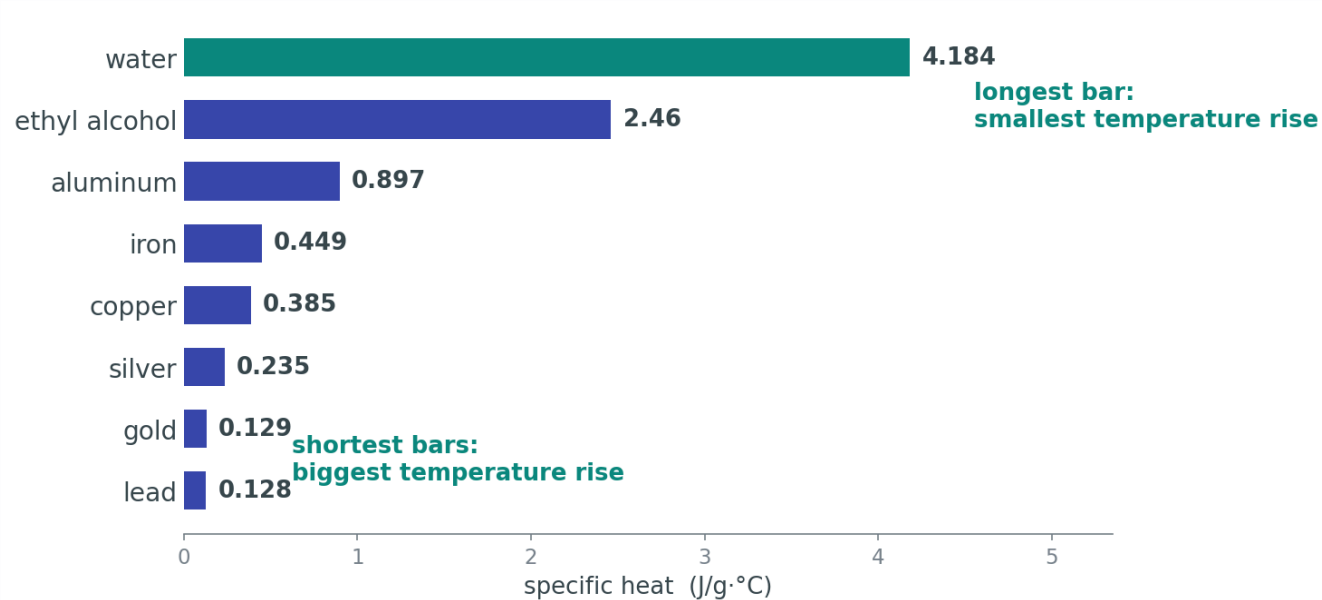


Give equal masses the same heat. Their temperature changes are not equal. Specific heat sits in the denominator of $\Delta T = q/mc$, so the smaller c , the larger the temperature change.

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

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.



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 $+26.0^\circ\text{C}$ · aluminum $+44.6^\circ\text{C}$ · 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

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

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 · 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

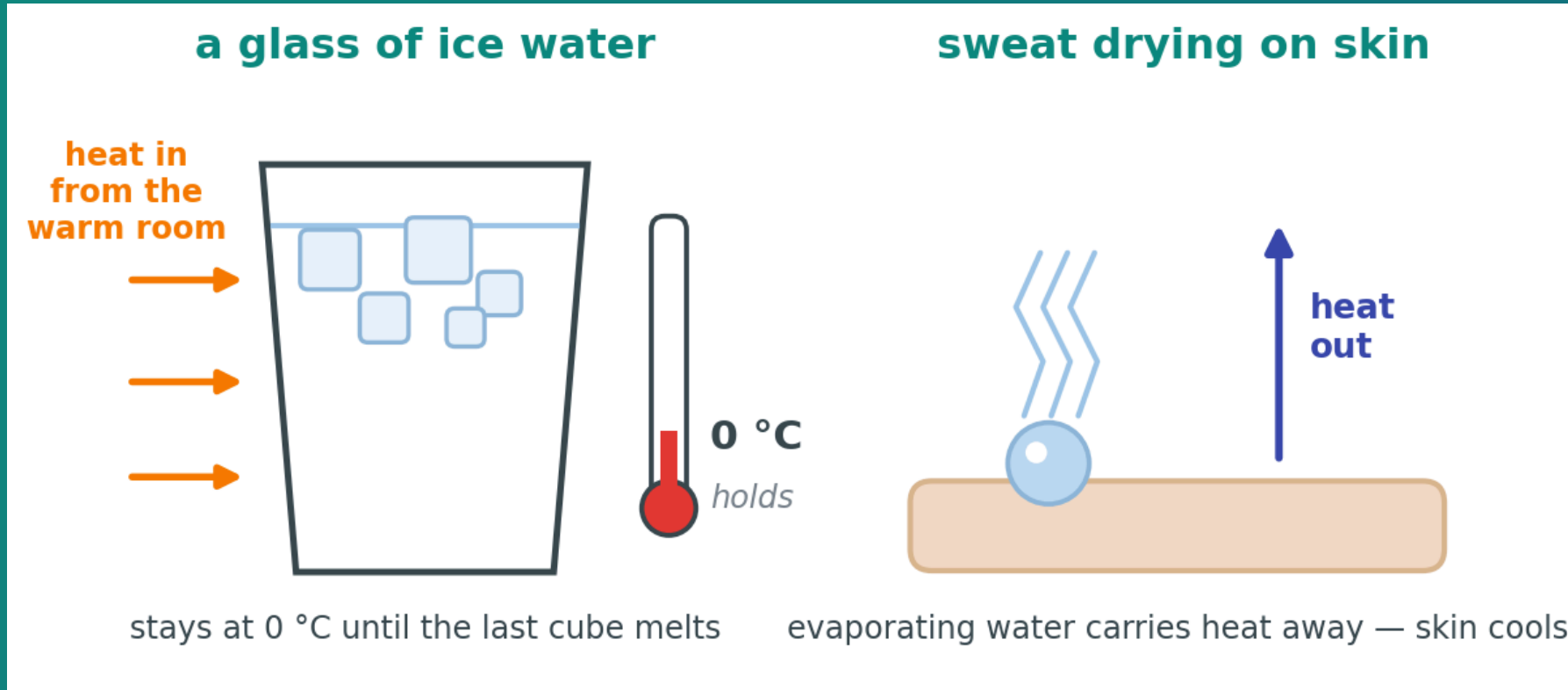
1. Equal masses of copper and water absorb the same heat from the same start. Which ends hotter, and which term in $\Delta T = q/mc$ decides it?
2. Two iron bars, 50 g and 150 g, absorb the same heat. Which shows the larger temperature change, and why?

Heat can also flow in while the thermometer holds still. At 0 °C ice melts, and at 100 °C water boils, with no temperature change at all. That heat goes into changing the phase, and it has its own per-gram price.

6 • Heat of a Phase Change

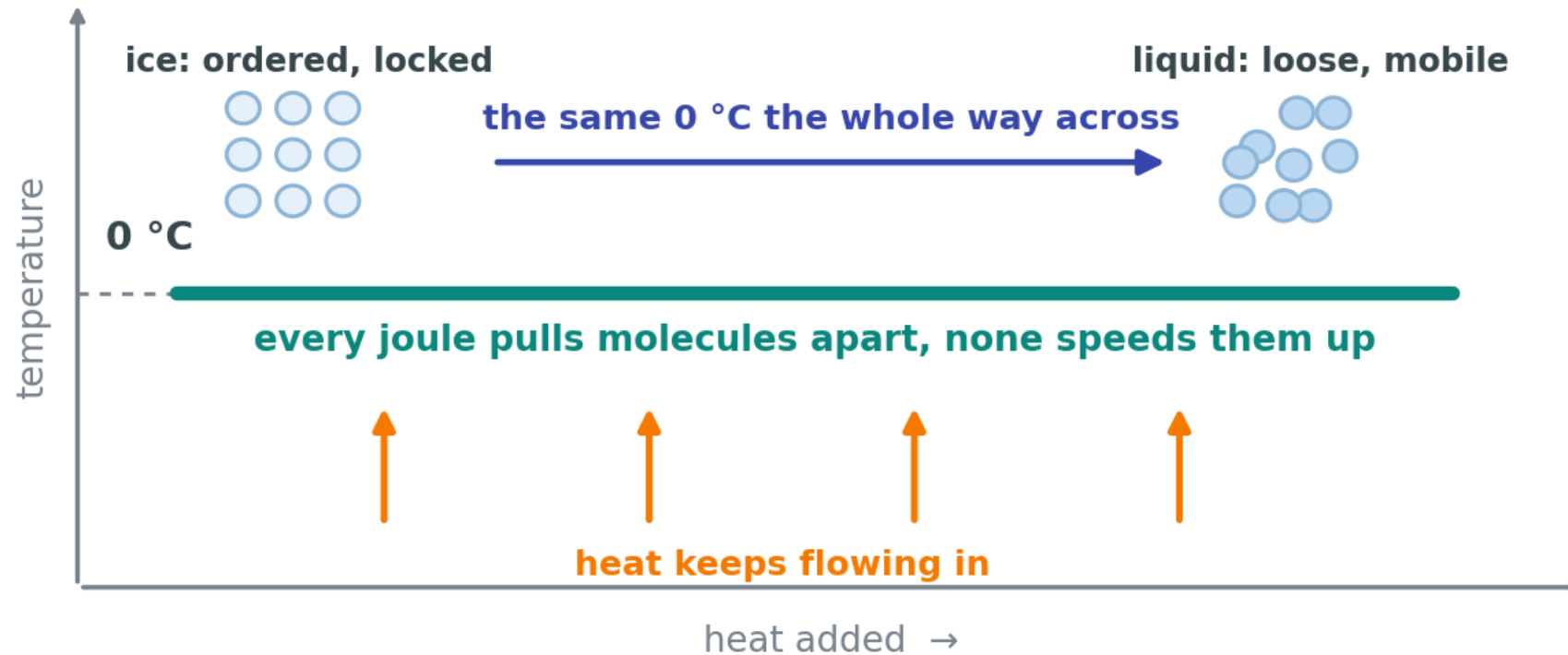
Find the heat of a phase change with $q = \text{mass} \times \Delta H$ in J/g or $q = n \times \Delta H$ in kJ/mol, choosing the heat of fusion for melting or freezing and the heat of vaporization for boiling or condensing, while the temperature holds constant.

Heat flows, the temperature holds



A glass of ice water stays at 0 °C until the last cube melts. Sweat cools your skin. Heat moves in or out; the temperature holds.

Temperature holds during a phase change



A phase change runs at one temperature. The heat does not warm the sample; it pulls the molecules apart. Ice water holds at 0 °C until the last cube melts.

Six changes of state

absorbs heat · endothermic



releases heat · exothermic



Changes toward the gas pull particles apart and absorb heat: endothermic. The reverses release that same heat: exothermic. Vaporization is also called evaporation. Dry ice goes straight from solid to gas; so does iodine warmed gently.

Heat of fusion and heat of vaporization

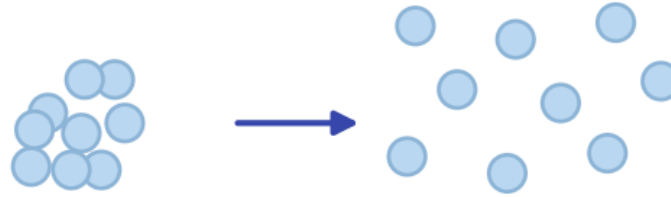
**for water: $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · $\Delta H_{\text{vap}} = 2259 \text{ J/g}$
melt or freeze: 335 J per gram · boil or condense: 2259 J per gram**

Melting is called fusion. The heat of fusion, ΔH_{fus} , melts one gram; the heat of vaporization, ΔH_{vap} , boils one gram. Freezing and condensing release the same amounts, reversed.

Vaporizing costs more than melting



melting: loosen the packing
(molecules still touch)



vaporizing: pull fully apart
(molecules fly free)

melt  **heat of fusion 335 J/g**

vaporize  **heat of vaporization 2259 J/g**

water: vaporizing takes about 7 times the heat of melting

ΔH_{vap} is far larger than ΔH_{fus} . Melting only loosens the packing; the molecules still touch. Vaporizing pulls them fully apart, which takes about seven times the energy.

The phase-change heat equation

$$q = \text{mass} \times \Delta H$$

$$\text{g} \times (\text{J/g}) = \text{J} \cdot \text{absorbed (+) to melt or boil} \cdot \text{released (-) to freeze or condense}$$

Each heat is a conversion factor in joules per gram. Multiply by the mass and grams cancel. Melting and boiling absorb heat; freezing and condensing release it, so q turns negative.

The same heats, per mole

$$q = n \times \Delta H$$

for water: $\Delta H_{\text{fus}} = 6.02 \text{ kJ/mol}$ · $\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$ · grams → moles → kilojoules

Phase-change heats also come per mole: 6.02 kJ melts a mole of ice, 40.7 kJ boils a mole of water. A mass in grams becomes moles first; either route lands on the same heat.



The method

1. **Name the phase change.**
2. **Pick its ΔH :** fusion to melt or freeze, vaporization to boil or condense.
3. **Multiply:** mass $\times \Delta H$ (J/g), or moles $\times \Delta H$ (kJ/mol).
4. **Set the direction:** melting and boiling absorb; freezing and condensing release.



Worked example 1: melt ice

$$q = \text{mass} \times \Delta H_{\text{fus}}$$

given: 60.0 g ice at 0 °C · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · wanted: q

A tray holds **60.0 g of ice** at 0 °C. How much heat melts it completely? (ΔH_{fus} of water: 335 J/g)

Name the change of state, then pick its heat.

✓ Worked example 1: solution

$$q = \text{mass} \times \Delta H_{\text{fus}}$$

given: 60.0 g ice at 0 °C · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · wanted: q

Step 1 · Name the phase change

Step 2 · Pick its ΔH

Step 3 · Multiply

$$60.0 \text{ g} \times \frac{335 \text{ J}}{1 \text{ g}} = 20,100 \text{ J}$$

Step 4 · Set the direction

 **Sense check:** Melting absorbs heat, so q is positive: 20,100 J (20.1 kJ) go in, and the temperature never leaves 0 °C. ✓

✓ Worked example 1: the route on the map

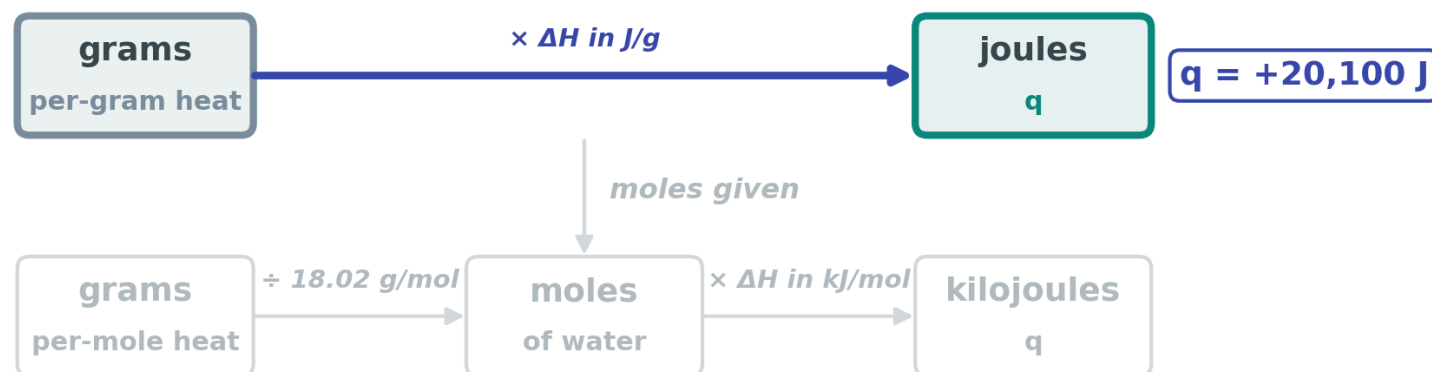
$$q = \text{mass} \times \Delta H_{\text{fus}}$$

given: 60.0 g ice at 0 °C · found: $q = +20,100 \text{ J}$

Steps 1, 2 and 4 · which ΔH , which sign



Step 3 · the route



🧠 **Sense check:** Melting picks ΔH_{fus} and a plus sign. Grams in with a per-gram heat: the top lane, one factor. ✓



Worked example 2: boil water to steam

$$q = \text{mass} \times \Delta H_{\text{vap}}$$

given: 20.0 g water at 100 °C · $\Delta H_{\text{vap}} = 2259 \text{ J/g}$ · wanted: q

20.0 g of water at 100 °C boils away to steam. How much heat does it take? (ΔH_{vap} of water: 2259 J/g)

Same route as melting, with the vaporization heat.

✓ Worked example 2: solution

$$q = \text{mass} \times \Delta H_{\text{vap}}$$

given: 20.0 g water at 100 °C · $\Delta H_{\text{vap}} = 2259 \text{ J/g}$ · wanted: q

Step 1 · Name the phase change

Step 2 · Pick its ΔH

Step 3 · Multiply

$$20.0 \text{ g} \times \frac{2259 \text{ J}}{1 \text{ g}} = 45,180 \text{ J}$$

Step 4 · Set the direction

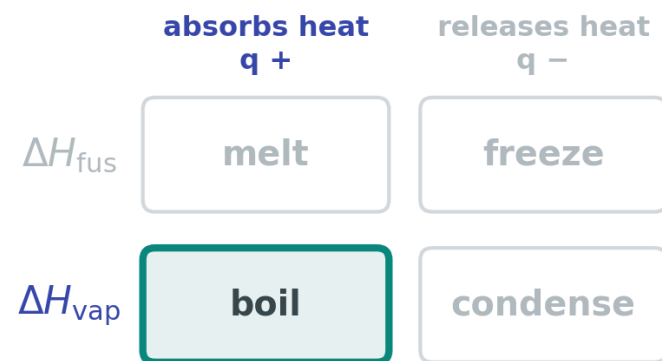
 **Sense check:** Boiling absorbs heat: $q = +45,180 \text{ J}$ (45.2 kJ). Same 20.0 g would take only 6,700 J to melt: vaporizing costs far more. ✓

✓ Worked example 2: the route on the map

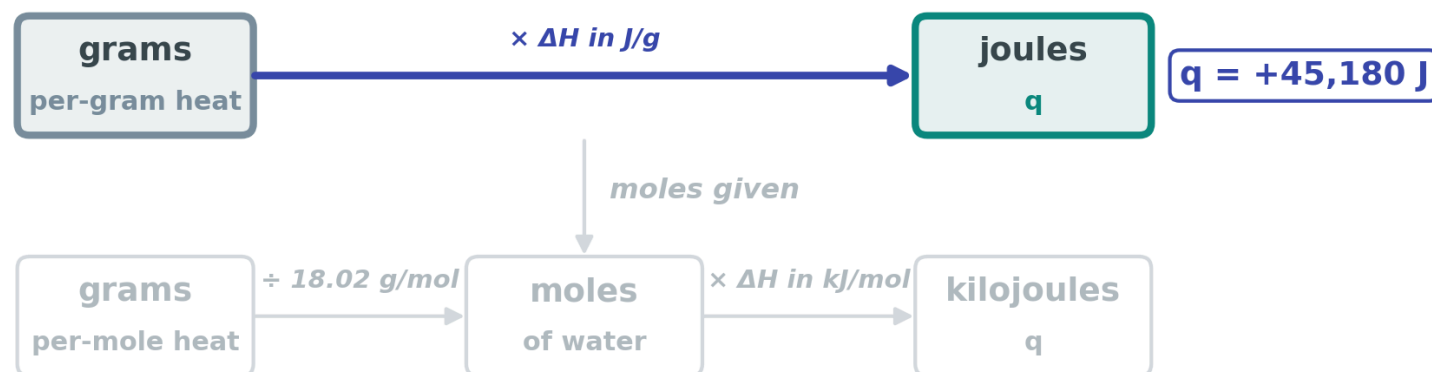
$$q = \text{mass} \times \Delta H_{\text{vap}}$$

given: 20.0 g water at 100 °C · found: $q = +45,180 \text{ J}$

Steps 1, 2 and 4 · which ΔH , which sign



Step 3 · the route



🧠 **Sense check:** Boiling picks ΔH_{vap} . Same top lane as melting; only the factor changed. ✓

Your turn: melt more ice

$$q = \text{mass} \times \Delta H_{\text{fus}}$$

given: 40.0 g ice at 0 °C · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · wanted: q

40.0 g of ice at 0 °C melts to water. Fill in the heat of fusion, then compute.

$$40.0 \text{ g} \times \frac{\text{ } \text{J}}{1 \text{ g}} = \text{ } \text{J}$$

$$40.0 \text{ g} \times \frac{335 \text{ J}}{1 \text{ g}} = \mathbf{13,400 \text{ J}}$$

 **Sense check:** 335 J melts one gram, so 40.0 g take $40.0 \times 335 = 13,400 \text{ J}$ (13.4 kJ), all at 0 °C. ✓

⚠ Where this goes wrong

$$q = \text{mass} \times \Delta H_{\text{fus}}$$

$$60.0 \text{ g ice at } 0 \text{ }^{\circ}\text{C} \cdot \Delta H_{\text{fus}} = 335 \text{ J/g} \cdot \text{correct } q = 20,100 \text{ J}$$

Leaving out the mass. 335 J melts a single gram. The sample has 60.0 of them. Scale it up: $60.0 \text{ g} \times 335 \text{ J/g} = 20,100 \text{ J}$.

Dividing by the heat of fusion. $60.0 \div 335 = 0.179$, in units of g^2/J . Nothing cancels. Write the factor so grams cancel: $60.0 \text{ g} \times (335 \text{ J} / 1 \text{ g})$.

Reporting kilojoules as joules. $60.0 \times 335 = 20,100$, then sliding the decimal gives 20.1. That is the value in kilojoules. In joules it is 20,100 J.

Using the vaporization heat to melt. Melting uses $\Delta H_{\text{fus}} = 335 \text{ J/g}$, not $\Delta H_{\text{vap}} = 2259 \text{ J/g}$. $60.0 \times 2259 = 135,540 \text{ J}$ is the heat to boil the water, not melt the ice.

 Practice 1

$$q = \text{mass} \times \Delta H_{\text{vap}}$$

given: $5.00 \times 10^4 \text{ J}$ absorbed by water at 100°C · $\Delta H_{\text{vap}} = 2259 \text{ J/g}$ · wanted: mass boiled away

A kettle delivers $5.00 \times 10^4 \text{ J}$ to water already boiling at 100°C . What mass of water, in grams, boils away? (ΔH_{vap} of water: 2259 J/g)

- A. 22.1
- B. 149
- C. 1.20×10^4
- D. 1.13×10^8

✓ Practice 1 answer: A

$$q = \text{mass} \times \Delta H_{\text{vap}}$$

given: $5.00 \times 10^4 \text{ J}$ absorbed at 100°C · $\Delta H_{\text{vap}} = 2259 \text{ J/g}$ · wanted: g

$$5.00 \times 10^4 \text{ J} \times \frac{1 \text{ g}}{2259 \text{ J}} = 22.1 \text{ g} \cdot \text{answer A}$$

The heat is given, so ΔH_{vap} enters upside down: J on the bottom, so joules cancel. B used the heat of fusion: $5.00 \times 10^4 \div 335 = 149 \text{ g}$ is the ice this heat would melt, not the water it boils. C used water's specific heat: $5.00 \times 10^4 \div 4.184 = 1.20 \times 10^4$, but a plateau has no ΔT , so c does not apply. D kept the factor right side up: $5.00 \times 10^4 \times 2259 = 1.13 \times 10^8$, in J^2/g , nothing cancels.

 **Sense check:** Each gram takes 2259 J to boil, so 50,000 J boils a little over 20 g. ✓



Worked example 3: a steam burn

$$q = \text{mass} \times \Delta H_{\text{vap}}$$

given: 8.00 g steam at 100 °C → water at 100 °C · $\Delta H_{\text{vap}} = 2259 \text{ J/g}$ · wanted: heat released

8.00 g of steam at 100 °C condenses on skin and releases heat. This is why a steam burn is so severe. How much heat comes out?

A common first attempt: add a $q = \text{mass} \times c \times \Delta T$ term for the temperature. Test it.

✓ Worked example 3: solution

$$q = \text{mass} \times \Delta H_{\text{vap}}$$

8.00 g steam · condenses at 100 °C → water at 100 °C · wanted: heat released

A common first attempt

$$m \cdot c \cdot \Delta T = 8.00 \text{ g} \times 4.184 \text{ J/g} \cdot ^\circ\text{C} \times (100 - 100) ^\circ\text{C} = \mathbf{0 \text{ J}}$$

The steam condenses at 100 °C into water at 100 °C. $\Delta T = 0$, so the $m \cdot c \cdot \Delta T$ term adds nothing.

Step 1 · Name the phase change

Step 2 · Pick its ΔH

Condensing uses the heat of vaporization, $\Delta H_{\text{vap}} = 2259 \text{ J/g}$, the same value as boiling.

 **Sense check:** No temperature change means no $m \cdot c \cdot \Delta T$ term. The phase-change heat is the whole answer. ✓

✓ Worked example 3: heat released

$$q = \text{mass} \times \Delta H_{\text{vap}}$$

$$8.00 \text{ g steam condensing at } 100 \text{ }^{\circ}\text{C} \cdot \Delta H_{\text{vap}} = 2259 \text{ J/g}$$

Step 3 • Multiply

$$8.00 \text{ g} \times \frac{2259 \text{ J}}{1 \text{ g}} = 18,072 \text{ J}$$

Step 4 • Set the direction

Heat leaves the steam as it condenses, so q is negative: $q = -18,072 \text{ J}$.

 **Sense check:** Condensing just 8.00 g of steam dumps 18,072 J (18.1 kJ) into the skin, all at 100 °C before the water even starts to cool. That is why steam burns are severe. ✓

✓ Worked example 3: the route on the map

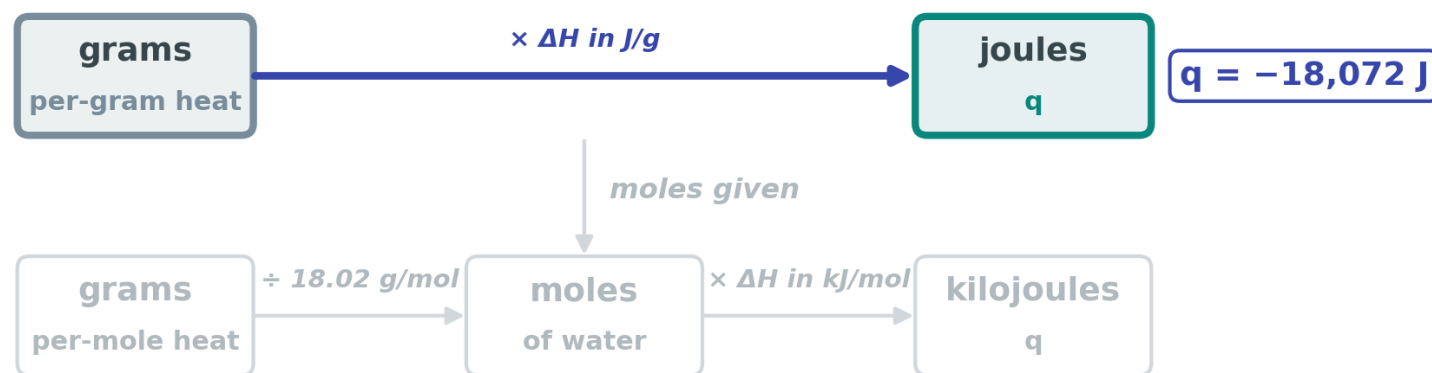
$$q = \text{mass} \times \Delta H_{\text{vap}}$$

given: 8.00 g steam condenses at 100 °C · found: $q = -18,072 \text{ J}$

Steps 1, 2 and 4 · which ΔH , which sign

	absorbs heat $q +$	releases heat $q -$
ΔH_{fus}	melt	freeze
ΔH_{vap}	boil	condense

Step 3 · the route



Sense check: Condensing sits in the releases column: ΔH_{vap} with a minus sign. The route is the top lane, with no $m \cdot c \cdot \Delta T$ anywhere. ✓



Worked example 4: melting by the molar route

$$q = n \times \Delta H_{\text{fus}}$$

given: 60.0 g ice at 0 °C · $\Delta H_{\text{fus}} = 6.02 \text{ kJ/mol}$ · molar mass of water 18.02 g/mol · wanted: q in kJ

60.0 g of ice at 0 °C melts to water. Find the heat with the molar heat of fusion, **6.02 kJ/mol**. The heat is per mole, so the grams must become moles first.

✓ Worked example 4: solution

$$q = n \times \Delta H_{\text{fus}}$$

given: 60.0 g ice at 0 °C · $\Delta H_{\text{fus}} = 6.02 \text{ kJ/mol}$ · molar mass of water 18.02 g/mol

Step 1 · Name the phase change

Step 2 · Pick its ΔH

Step 3 · Multiply

$$60.0 \text{ g} \times \frac{1 \text{ mol}}{18.02 \text{ g}} \times \frac{6.02 \text{ kJ}}{1 \text{ mol}} = 20.0 \text{ kJ}$$

Step 4 · Set the direction

The per-gram route on the same sample: $60.0 \text{ g} \times 335 \text{ J/g} = 20,100 \text{ J} = 20.1 \text{ kJ}$. Two routes, one heat, agreeing to the rounding of the constants.

 **Sense check:** Melting absorbs heat: $q = +20.0 \text{ kJ}$. 3.33 mol of ice at 6.02 kJ each is about 20 kJ, and the 335 J/g route lands on the same number. ✓

✓ Worked example 4: the route on the map

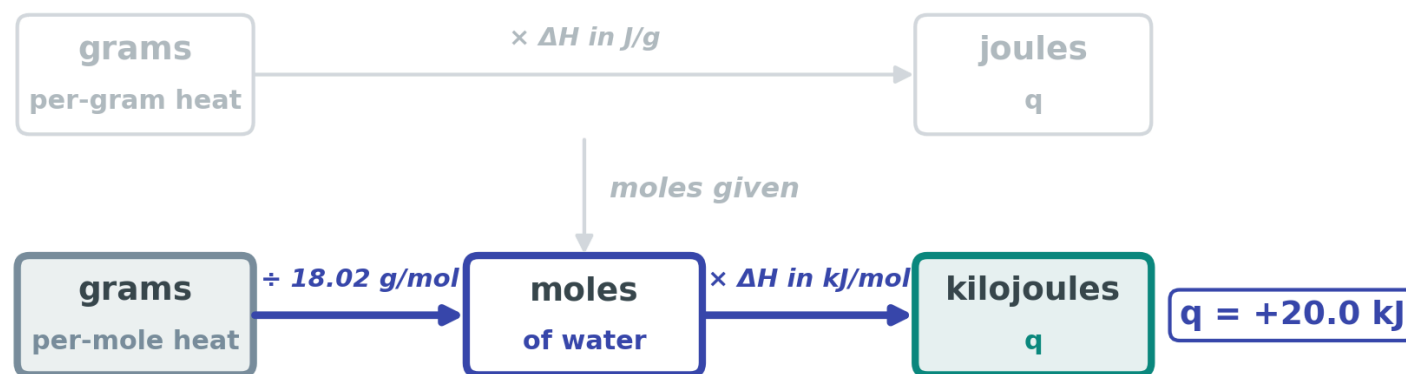
$$q = n \times \Delta H_{\text{fus}}$$

given: 60.0 g ice at 0 °C · found: $q = +20.0 \text{ kJ}$

Steps 1, 2 and 4 · which ΔH , which sign



Step 3 · the route



 **Sense check:** Grams given with a heat per mole: the bottom lane, molar mass first, then ΔH_{fus} . Two factors. ✓

Take-home: a phase change has no ΔT

temperature changes, one phase: $q = m \cdot c \cdot \Delta T$
warming or cooling within a solid, liquid, or gas

temperature constant, changing phase: $q = \text{mass} \times \Delta H$ or $q = n \times \Delta H$
melting, freezing, boiling, condensing. $\Delta T = 0$, so no $m \cdot c \cdot \Delta T$ term

During a phase change the temperature holds, so $\Delta T = 0$ and the $m \cdot c \cdot \Delta T$ term is zero. Use ΔH alone. The two never combine within one phase change.

 Practice 2

$$q = n \times \Delta H$$

given: 15.0 kJ removed from liquid water at 0 °C · $\Delta H_{\text{fus}} = 6.02 \text{ kJ/mol}$ · $\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$ · water
18.02 g/mol · wanted: g of ice

A freezer removes **15.0 kJ** of heat from liquid water at 0 °C. What mass of ice, in grams, forms?

- A. 6.64
- B. 44.9
- C. 0.138
- D. 2.49

✓ Practice 2 answer: B

$$q = n \times \Delta H_{\text{fus}}$$

freezing is melting reversed: the fusion heat applies · 15.0 kJ released by the water · 18.02 g/mol

$$15.0 \text{ kJ} \times \frac{1 \text{ mol}}{6.02 \text{ kJ}} \times \frac{18.02 \text{ g}}{1 \text{ mol}} = 44.9 \text{ g} \cdot \text{answer B}$$

The heat is given, so ΔH_{fus} enters upside down and kJ cancels. A picked the vaporization heat: $15.0 \div 40.7 \times 18.02 = 6.64 \text{ g}$, but no boiling or condensing happens here. C divided by the molar mass: $2.49 \div 18.02 = 0.138 \text{ g}$. D stopped at moles: 2.49 mol of water, one factor short of grams.

 **Sense check:** Freezing releases heat, so q for the water is -15.0 kJ ; the mass of ice is positive. 2.49 mol at 18.02 g each is about 45 g. ✓

 Practice 3

$$q = \text{mass} \times \Delta H$$

given: 28.0 g boiled vs 28.0 g melted · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · $\Delta H_{\text{vap}} = 2259 \text{ J/g}$ · wanted: difference

How much more heat, in joules, does it take to boil away **28.0 g** of water at 100 °C than to melt 28.0 g of ice at 0 °C?

- A. 7.26×10^4
- B. 6.74
- C. 6.33×10^4
- D. 5.39×10^4

✓ Practice 3 answer: D

$q = \text{mass} \times \Delta H$, once for each change

28.0 g · boil: $\Delta H_{\text{vap}} = 2259 \text{ J/g}$ · melt: $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · wanted: boil minus melt

$$28.0 \text{ g} \times \frac{2259 \text{ J}}{1 \text{ g}} = 63,252 \text{ J} \quad \cdot \quad 28.0 \text{ g} \times \frac{335 \text{ J}}{1 \text{ g}} = 9,380 \text{ J}$$

$$63,252 \text{ J} - 9,380 \text{ J} = 5.39 \times 10^4 \text{ J} \cdot \text{answer D}$$

A added the two heats: $63,252 + 9,380 = 7.26 \times 10^4 \text{ J}$ is the cost of both changes, not the gap. B divided the heats: $2259 \div 335 = 6.74$ says how many times more, not how many joules more. C stopped at the boiling heat: $28.0 \times 2259 = 6.33 \times 10^4 \text{ J}$, with the melt never subtracted.

 **Sense check:** Each gram costs $2259 - 335 = 1924 \text{ J}$ more to boil than to melt. 28.0 g of that is about 54,000 J. ✓



Practice 4

$$q = n \times \Delta H$$

given: 62.0 g C₂H₅OH boiled · ΔH_{vap} of ethanol = 38.6 kJ/mol · wanted: q in kJ

A still boils off **62.0 g** of ethanol, C₂H₅OH, at its boiling point. How much heat, in kJ, does the ethanol absorb?

- A. 51.9
- B. 2.39×10^3
- C. 133
- D. 1.10×10^5

✓ Practice 4 answer: A

$$q = n \times \Delta H_{\text{vap}}$$

boiling: the vaporization heat applies · 62.0 g C₂H₅OH · 38.6 kJ/mol · molar mass from the formula

Two conversion factors are needed, and the first one is ethanol's molar mass: $2(12.01) + 6(1.008) + 16.00 = 46.07 \text{ g/mol}$.

$$62.0 \text{ g} \times \frac{1 \text{ mol}}{46.07 \text{ g}} \times \frac{38.6 \text{ kJ}}{1 \text{ mol}} = 51.9 \text{ kJ} \cdot \text{answer A}$$

B skipped the molar mass: $62.0 \times 38.6 = 2.39 \times 10^3$, as if each gram were a mole. C used water's molar mass: $62.0 \div 18.02 \times 38.6 = 133$, but the liquid is ethanol. D put the molar mass upside down: $62.0 \times 46.07 \times 38.6 = 1.10 \times 10^5$, and grams do not cancel.

 **Sense check:** Boiling absorbs heat: $q = +51.9 \text{ kJ}$. About 1.35 mol at about 39 kJ each is about 52 kJ. ✓

 Practice 5

$$q = \text{mass} \times \Delta H_{\text{fus}}, \text{ then } q = m \cdot c \cdot \Delta T$$

given: 32.0 g ice melts · the same heat instead goes into 400. g water at 21.0 °C · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · $c = 4.184 \text{ J/g}\cdot^\circ\text{C}$

Melting **32.0 g** of ice at 0 °C takes a certain heat. If that same heat instead warmed 400. g of water at **21.0 °C**, what final temperature, in °C, would the water reach?

- A. 6.41
- B. 101.1
- C. 27.4
- D. 64.2

✓ Practice 5 answer: C

$$q = \text{mass} \times \Delta H_{\text{fus}}, \text{ then } \Delta T = q \div (m \cdot c)$$

32.0 g ice melts · the same heat instead goes into 400. g water at 21.0 °C · 335 J/g · 4.184 J/g·°C

$$32.0 \text{ g} \times \frac{335 \text{ J}}{1 \text{ g}} = 10,720 \text{ J} \rightarrow \Delta T = \frac{10,720 \text{ J}}{400. \text{ g} \times 4.184 \text{ J/g} \cdot ^\circ\text{C}} = 6.41 ^\circ\text{C}$$

$$T_{\text{final}} = 21.0 ^\circ\text{C} + 6.41 ^\circ\text{C} = \mathbf{27.4 ^\circ\text{C} \cdot \text{answer C}}$$

A stopped at ΔT : 6.41 °C is the rise, not the final temperature. B warmed the ice's 32.0 g: $21.0 + 10,720 \div (32.0 \times 4.184) = 101.1$, but the 400. g of water takes the heat. D used the heat of vaporization: $32.0 \times 2259 = 72,288 \text{ J}$, giving 64.2, but nothing boils.

 **Sense check:** The water took in heat instead of the ice, so 27.4 °C sits above 21.0 °C. ✓

Check yourself

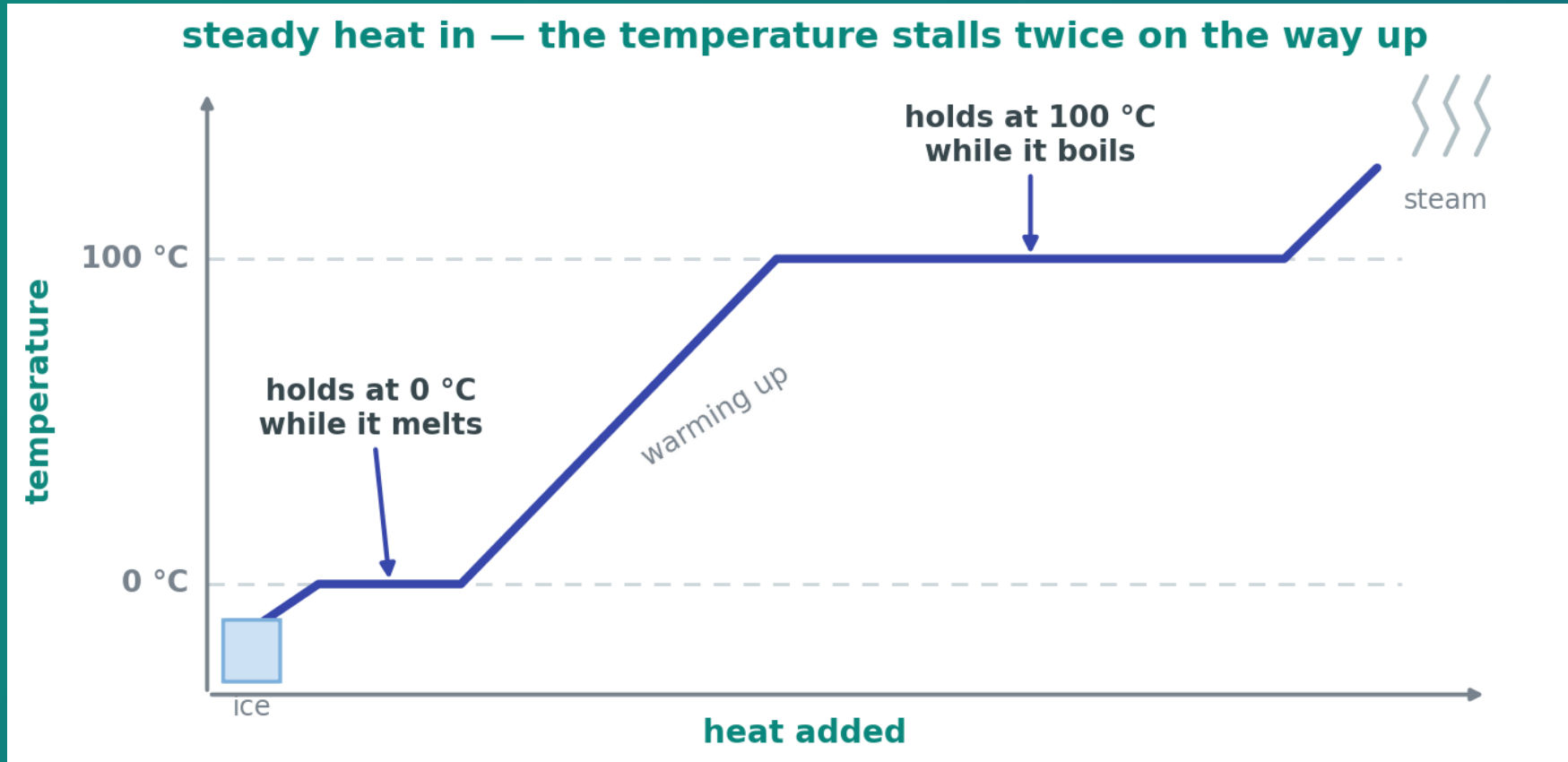
1. A block of ice at 0 °C melts to water at 0 °C. What happens to the temperature while it melts, and where does the heat go?
2. Which is larger for water, the heat of fusion or the heat of vaporization, and why?

Melting and boiling are the flat steps of a heating curve. Between those steps the temperature climbs, and there $q = m \cdot c \cdot \Delta T$ takes over. The full curve chains both kinds of heat, segment by segment.

7 · Heating & Cooling Curves

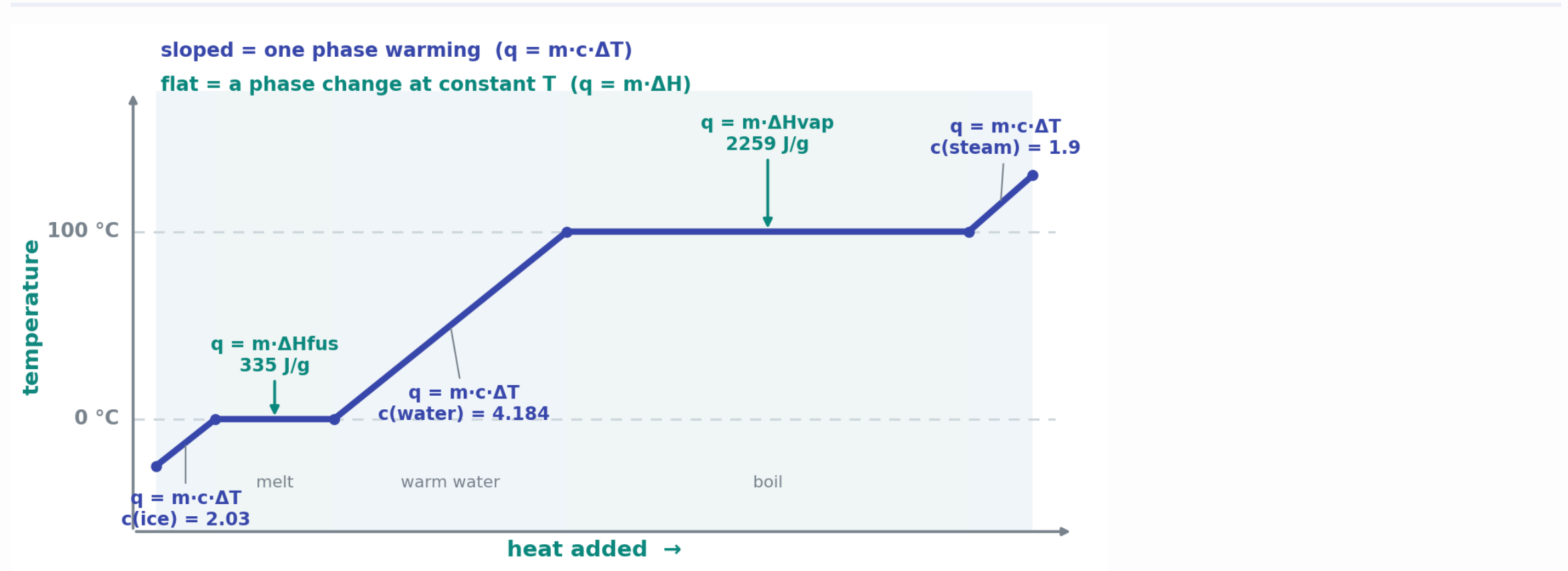
Break a heating path into segments, use $q = m \cdot c \cdot \Delta T$ for each slope and the phase-change energy for each plateau, and add them for the total heat.

From the freezer to a rolling boil



Heat a block of ice steadily. Temperature climbs, holds at 0 °C while it melts, climbs, then holds at 100 °C while it boils.

Warming and phase changes alternate



Heating a substance alternates between warming one phase (a sloped step, $q = m \cdot c \cdot \Delta T$) and changing the phase (a flat step, phase-change energy). The total heat is the sum of every step.

Two kinds of segment, two equations

sloped segment: one phase warming

$q = m \cdot c \cdot \Delta T$ · uses that phase's specific heat c

flat plateau: a phase change at constant T

$q = m \cdot \Delta H$ · the temperature does not move, so there is no ΔT

Read the graph one segment at a time. A slope warms a single phase. A plateau holds the temperature fixed while the phase changes. Each segment needs its own equation.

Water's constants for each segment

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

$$c(\text{ice}) = 2.03 \cdot c(\text{liquid water}) = 4.184 \cdot c(\text{steam}) = 1.9 \text{ J/g}\cdot^\circ\text{C}$$

plateaus: $q = m \cdot \Delta H$

$$\Delta H_{\text{fus}} = 335 \text{ J/g at } 0^\circ\text{C} \cdot \Delta H_{\text{vap}} = 2259 \text{ J/g at } 100^\circ\text{C}$$

Each phase carries its own specific heat. Each phase change carries its own energy. Ice and steam sit near 2 J/g·°C; only liquid water earns 4.184.



The method

1. **Identify each segment.** Each slope and each plateau is one piece.
2. **Compute each piece.** A slope uses $q = m \cdot c \cdot \Delta T$ with that phase's c . A plateau uses the phase-change energy, ΔH_{fus} or ΔH_{vap} .
3. **Add every piece** for the total.



Worked example 1: ice at 0 °C to warm water

$$\text{total } q = \text{melt} + \text{warm}$$

given: 30 g ice at 0 °C → liquid water at 25 °C · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · $c(\text{water}) = 4.184 \text{ J/g}\cdot^\circ\text{C}$ · wanted: q

A **30 g** block of ice, already at 0 °C, is heated until it becomes liquid water at 25 °C. How much heat does it take?

Identify each segment, then add the pieces.

✓ Worked example 1: solution

total q = melt + warm

given: 30 g ice at 0 °C → liquid water at 25 °C · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · $c(\text{water}) = 4.184 \text{ J/g}\cdot^{\circ}\text{C}$

Step 1 · Identify each segment

Step 2 · Compute each piece

$$\text{melt: } 30 \text{ g} \times \frac{335 \text{ J}}{1 \text{ g}} = 10,050 \text{ J} \cdot \text{warm: } 30 \text{ g} \times \frac{4.184 \text{ J}}{1 \text{ g}\cdot^{\circ}\text{C}} \times 25 \text{ }^{\circ}\text{C} = 3138 \text{ J}$$

Step 3 · Add every piece

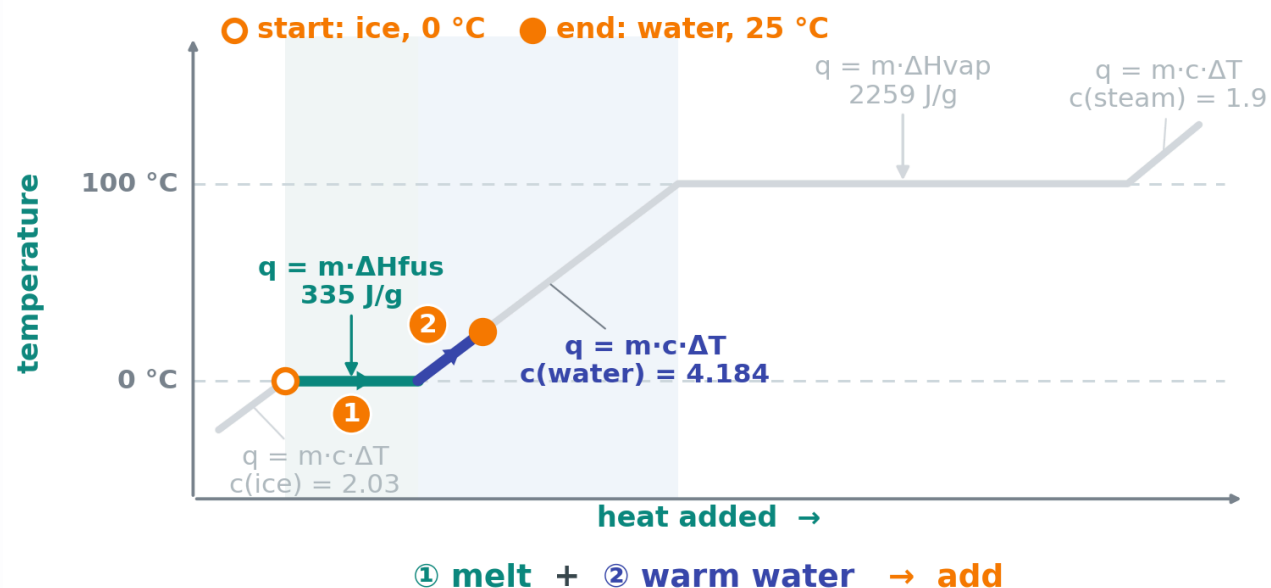
$$q = 10,050 \text{ J} + 3138 \text{ J} = 13,188 \text{ J}$$

 **Sense check:** Melting alone costs 10,050 J, more than warming the liquid 25 °C. The flat step costs more. ✓

✓ Worked example 1: the route on the curve

total $q = \text{melt} + \text{warm}$

given: 30 g ice at 0 °C → liquid water at 25 °C · found: $q = 13,188 \text{ J}$



🧠 **Sense check:** The ice starts on the plateau, so no ice warms. Two pieces: melt, then warm water. ✓



Worked example 2: melt, then warm to 40 °C

$$\text{total } q = \text{melt} + \text{warm}$$

given: 45 g ice at 0 °C → liquid water at 40 °C · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · $c(\text{water}) = 4.184 \text{ J/g}\cdot^\circ\text{C}$ · wanted: q

A **45 g** block of ice at 0 °C is melted and then warmed to 40 °C.

A common first attempt: the sample ends 40 °C warmer, so multiply the melting heat by 40 as well. Test it.

✓ Worked example 2: the melting step

total q = melt + warm

given: 45 g ice at 0 °C → liquid water at 40 °C · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · $c(\text{water}) = 4.184 \text{ J/g}\cdot^\circ\text{C}$

A common first attempt

$$\text{melt: } 45 \text{ g} \times \frac{335 \text{ J}}{1 \text{ g}} \times 40^\circ\text{C} = 603,000 \text{ J} \times$$

Step 1 • Identify each segment

Two pieces: melt at 0 °C, then warm the liquid from 0 °C to 40 °C. Only the warming piece has a ΔT .

 **Sense check:** The melting step holds at 0 °C, so it carries no temperature change. Only the warming step does. ✓

✓ Worked example 2: the two pieces

total q = melt + warm

given: 45 g ice at 0 °C → liquid water at 40 °C · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · $c(\text{water}) = 4.184 \text{ J/g}\cdot^{\circ}\text{C}$

Step 2 · Compute each piece

$$\text{melt: } 45 \text{ g} \times \frac{335 \text{ J}}{1 \text{ g}} = 15,075 \text{ J} \cdot \text{warm: } 45 \text{ g} \times \frac{4.184 \text{ J}}{1 \text{ g}\cdot^{\circ}\text{C}} \times 40 \text{ }^{\circ}\text{C} = 7531.2 \text{ J}$$

Step 3 · Add every piece

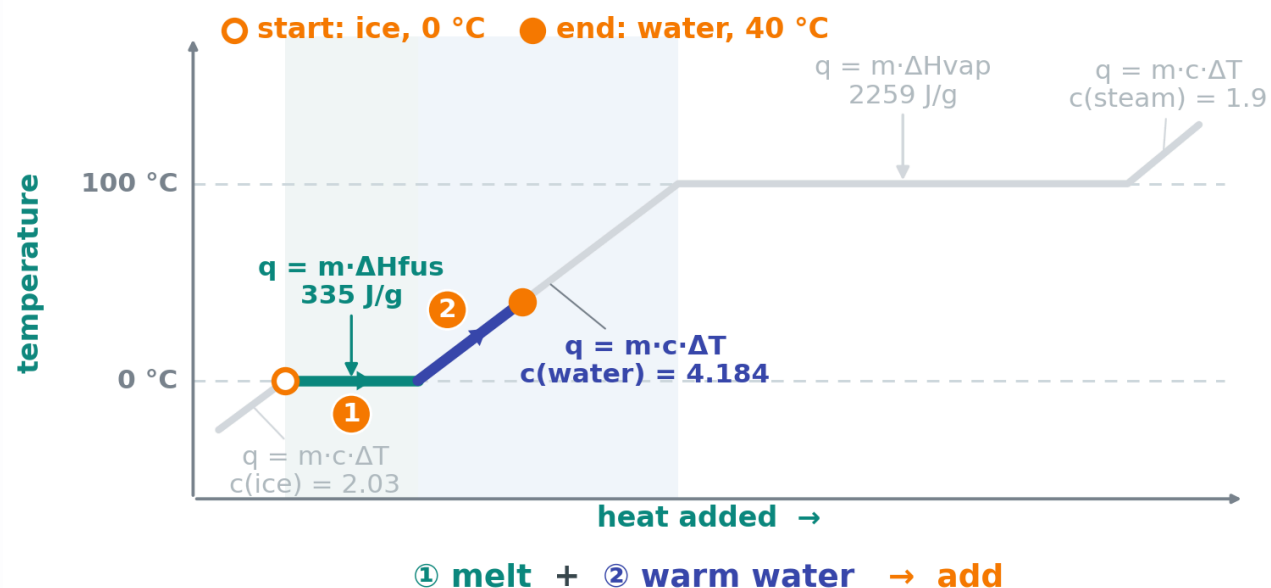
$$q = 15,075 \text{ J} + 7531.2 \text{ J} = 22,606 \text{ J}$$

 **Sense check:** The two pieces add to 22,606 J. The plateau-times- ΔT shortcut, 603,000 J, was about 27 times too large. ✓

✓ Worked example 2: the route on the curve

total $q = \text{melt} + \text{warm}$

given: 45 g ice at 0 °C → liquid water at 40 °C · found: $q = 22,606 \text{ J}$



🧠 **Sense check:** The same two pieces. Only the slope carries a ΔT ; the plateau is $m \cdot \Delta H_{\text{fus}}$ alone. ✓

Take-home: a plateau is not $q = m \cdot c \cdot \Delta T$

melt at 0 °C: correct

$$q = m \cdot \Delta H_{\text{fus}} = 45 \text{ g} \times 335 \text{ J/g} = 15,075 \text{ J} \cdot \text{temperature stays } 0 \text{ }^{\circ}\text{C}$$

treating the plateau as $q = m \cdot c \cdot \Delta T$: wrong

$$45 \text{ g} \times 335 \text{ J/g} \times 40 \text{ }^{\circ}\text{C} = 603,000 \text{ J} \times \cdot \text{there is no } \Delta T \text{ on a plateau}$$

On a flat plateau the temperature is constant, so ΔT is zero. The heat comes from ΔH_{fus} or ΔH_{vap} . A phase-change energy is never multiplied by a temperature change.



Worked example 3: ice at $-20\text{ }^{\circ}\text{C}$ to steam at $120\text{ }^{\circ}\text{C}$

total q = warm ice + melt + warm water + boil + warm steam

given: $20\text{ g} \cdot -20\text{ }^{\circ}\text{C} \rightarrow 120\text{ }^{\circ}\text{C} \cdot c(\text{ice}) 2.03 \cdot c(\text{water}) 4.184 \cdot c(\text{steam}) 1.9\text{ J/g}\cdot^{\circ}\text{C} \cdot \Delta H_{\text{fus}} 335 \cdot \Delta H_{\text{vap}} 2259\text{ J/g} \cdot$ wanted: q

A **20 g** sample starts as ice at $-20\text{ }^{\circ}\text{C}$ and ends as steam at $120\text{ }^{\circ}\text{C}$.

Count the segments between the start and end, then compute and add each one.

✓ Worked example 3: warming and melting

total q = warm ice + melt + warm water + boil + warm steam
 20 g · -20 °C → 120 °C · five segments to cross

Step 1 · Identify each segment

Step 2 · Compute each piece

$$\text{warm ice: } 20 \text{ g} \times \frac{2.03 \text{ J}}{1 \text{ g} \cdot ^\circ\text{C}} \times 20 ^\circ\text{C} = \mathbf{812 \text{ J}} \cdot \text{melt: } 20 \text{ g} \times \frac{335 \text{ J}}{1 \text{ g}} = \mathbf{6700 \text{ J}}$$

$$\text{warm water: } 20 \text{ g} \times \frac{4.184 \text{ J}}{1 \text{ g} \cdot ^\circ\text{C}} \times 100 ^\circ\text{C} = \mathbf{8368 \text{ J}}$$

🧠 **Sense check:** Each slope uses its own phase's c: ice 2.03, water 4.184. The plateaus still to come cost more. ✓

✓ Worked example 3: boiling, then the total

total q = warm ice + melt + warm water + boil + warm steam
 warm ice 812 · melt 6700 · warm water 8368 J so far

Step 2 · Compute each piece

$$\text{boil: } 20 \text{ g} \times \frac{2259 \text{ J}}{1 \text{ g}} = 45,180 \text{ J} \cdot \text{ warm steam: } 20 \text{ g} \times \frac{1.9 \text{ J}}{1 \text{ g} \cdot ^\circ\text{C}} \times 20 ^\circ\text{C} = 760 \text{ J}$$

Step 3 · Add every piece

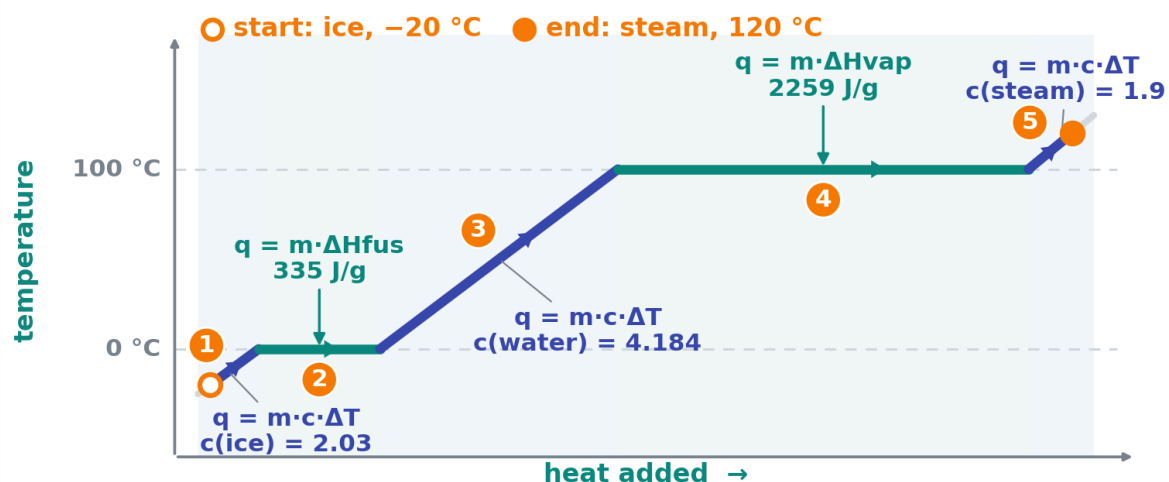
$$q = 812 + 6700 + 8368 + 45,180 + 760 = 61,820 \text{ J}$$

🧠 **Sense check:** Boiling alone is 45,180 J, about three-quarters of the total. Vaporizing water takes the most heat. ✓

✓ Worked example 3: the route on the curve

total q = warm ice + melt + warm water + boil + warm steam

given: 20 g ice at $-20\text{ }^{\circ}\text{C}$ $\rightarrow$ steam at $120\text{ }^{\circ}\text{C}$ · found: $q = 61,820\text{ J}$



① warm ice + ② melt + ③ warm water + ④ boil + ⑤ warm steam $\rightarrow$ add

🧠 **Sense check:** The whole curve: three slopes, each with its own c , and both plateaus. ✓

Your turn: 15 g of ice to warm water

total q = melt + warm

given: 15 g ice at 0 °C → liquid water at 50 °C · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · $c(\text{water}) = 4.184 \text{ J/g} \cdot ^\circ\text{C}$

melt: $15 \text{ g} \times (\text{ } \text{ J / 1 g}) = \text{ } \text{ J}$ · warm: $15 \text{ g} \times 4.184 \text{ J/(g} \cdot ^\circ\text{C)} \times \text{ } ^\circ\text{C} = 3138 \text{ J}$

Fill in ΔH_{fus} , the melt heat, and the temperature change, then add the two pieces.

melt: $15 \text{ g} \times (335 \text{ J / 1 g}) = \mathbf{5025 \text{ J}}$ · warm: $15 \text{ g} \times 4.184 \text{ J/(g} \cdot ^\circ\text{C)} \times 50 ^\circ\text{C} = \mathbf{3138 \text{ J}}$

total = $5025 \text{ J} + 3138 \text{ J} = \mathbf{8163 \text{ J}}$

 **Sense check:** The melt costs 5025 J, more than warming the liquid all the way to 50 °C. Melting is the larger step.



⚠ Where this goes wrong

reference: 30 g ice at 0 °C → liquid water at 25 °C

melt 10,050 J + warm 3138 J = 13,188 J · ΔH_{fus} 335 · c(ice) 2.03 · c(water) 4.184 J/g·°C

Treating a plateau as $q = m \cdot c \cdot \Delta T$. Multiplying the fusion heat by a ΔT , $45 \text{ g} \times 335 \text{ J/g} \times 40 \text{ °C} = 603,000 \text{ J}$, invents heat. Melting holds at 0 °C, so $45 \text{ g} \times 335 \text{ J/g} = 15,075 \text{ J}$, no ΔT .

Skipping the melting plateau. Warming the liquid only, $30 \text{ g} \times 4.184 \text{ J/g} \cdot \text{°C} \times 25 \text{ °C} = 3138 \text{ J}$, leaves the ice unmelted. Melting first costs another 10,050 J.

Stopping at the melting point. $30 \text{ g} \times 335 \text{ J/g} = 10,050 \text{ J}$ melts the ice but leaves it at 0 °C. Warming to 25 °C adds 3138 J.

Using the wrong phase's specific heat. Warming ice with water's 4.184, $20 \text{ g} \times 4.184 \times 20 \text{ °C} = 1673.6 \text{ J}$, overcharges it. Ice's c is 2.03: $20 \text{ g} \times 2.03 \times 20 \text{ °C} = 812 \text{ J}$.



Practice 1

total q = melt + warm

given: 20.0 g ice at 0 °C → liquid water at 30.0 °C · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · $c(\text{water}) = 4.184 \text{ J/g}\cdot^\circ\text{C}$ · wanted:
 q

A **20.0 g** block of ice at 0 °C is melted and warmed to 30.0 °C. How much heat, in J, does it take?

- A. 2.51×10^3
- B. 6.70×10^3
- C. 9.21×10^3
- D. 2.01×10^5

✓ Practice 1 answer: C

total q = melt + warm

given: 20.0 g ice at 0 °C → liquid water at 30.0 °C · $\Delta H_{\text{fus}} = 335 \text{ J/g}$ · $c(\text{water}) = 4.184 \text{ J/g}\cdot^{\circ}\text{C}$

$$\text{melt: } 20.0 \text{ g} \times (335 \text{ J} / 1 \text{ g}) = 6.70 \times 10^3 \text{ J} \cdot \text{warm: } 20.0 \text{ g} \times 4.184 \text{ J}/(\text{g} \cdot ^{\circ}\text{C}) \times 30.0 ^{\circ}\text{C} \\ = 2.51 \times 10^3 \text{ J}$$

$$q = 6.70 \times 10^3 \text{ J} + 2.51 \times 10^3 \text{ J} = \mathbf{9.21 \times 10^3 \text{ J} \cdot \text{answer C}}$$

A skipped the melting plateau: $20.0 \times 4.184 \times 30.0 = 2.51 \times 10^3 \text{ J}$ warms the liquid but never melts the ice. B stopped at the melting point: $20.0 \times 335 = 6.70 \times 10^3 \text{ J}$ leaves the water at 0 °C. D multiplied the plateau by a ΔT : $20.0 \times 335 \times 30.0 = 2.01 \times 10^5 \text{ J}$.

🧠 **Sense check:** Melting and warming are two separate costs, so both add ✓

 Practice 2

total q = cool steam + condense + cool water

given: 10.0 g steam at 115 °C → liquid water at 70.0 °C · $c(\text{steam})$ 1.9 · $c(\text{water})$ 4.184 J/g·°C · ΔH_{vap} 2259 J/g · wanted: q , signed

A **10.0 g** sample of steam at 115 °C cools, condenses, and ends as liquid water at 70.0 °C. What is q for the sample, in J?

- A. $+2.41 \times 10^4$
- B. -1.54×10^3
- C. -2.29×10^4
- D. -2.41×10^4

✓ Practice 2 answer: D

total q = cool steam + condense + cool water

given: $10.0 \text{ g} \cdot 115 \text{ }^{\circ}\text{C} \rightarrow 70.0 \text{ }^{\circ}\text{C} \cdot c(\text{steam}) 1.9 \cdot c(\text{water}) 4.184 \text{ J/g}\cdot^{\circ}\text{C} \cdot \Delta H_{\text{vap}} 2259 \text{ J/g} \cdot \Delta T = T_{\text{final}} - T_{\text{initial}}$

cool steam: $10.0 \text{ g} \times 1.9 \times (100 - 115) \text{ }^{\circ}\text{C} = -285 \text{ J} \cdot \text{condense: } -(10.0 \text{ g} \times 2259 \text{ J/g}) = -22,590 \text{ J} \cdot \text{cool water: } 10.0 \text{ g} \times 4.184 \times (70.0 - 100) \text{ }^{\circ}\text{C} = -1255 \text{ J}$

$q = -285 \text{ J} + (-22,590 \text{ J}) + (-1255 \text{ J}) = -24,130 \text{ J} = -2.41 \times 10^4 \text{ J} \cdot \text{answer D}$

A lost the sign: a cooling path releases heat, so q is negative. B skipped the condensation plateau: $-285 - 1255 = -1.54 \times 10^3 \text{ J}$. C stopped at $100 \text{ }^{\circ}\text{C}$: $-285 - 22,590 = -2.29 \times 10^4 \text{ J}$.

🧠 **Sense check:** Heating curve run backward: every step is negative, condensing dominates ✓



Practice 3

total q = cool water + freeze

given: 50.0 g liquid water at 22.0 °C → ice at 0 °C · $c(\text{water})$ 4.184 · $c(\text{ice})$ 2.03 J/g·°C · ΔH_{fus} 335 J/g ·
wanted: q , signed

A bottle holding **50.0 g** of water at **22.0 °C** goes into a freezer and ends as ice at 0 °C.
What is q for the water, in J?

- A. -1.68×10^4
- B. $+2.14 \times 10^4$
- C. -4.60×10^3
- D. -2.14×10^4
- E. -1.90×10^4

✓ Practice 3 answer: D

total q = cool water + freeze

given: $50.0 \text{ g} \cdot 22.0 \text{ }^{\circ}\text{C} \rightarrow \text{ice at } 0 \text{ }^{\circ}\text{C} \cdot c(\text{water}) 4.184 \text{ J/g}\cdot^{\circ}\text{C} \cdot \Delta H_{\text{fus}} 335 \text{ J/g} \cdot \Delta T = T_{\text{final}} - T_{\text{initial}}$

cool water: $50.0 \text{ g} \times 4.184 \text{ J/(g}\cdot^{\circ}\text{C)} \times (0 - 22.0) \text{ }^{\circ}\text{C} = -4602 \text{ J}$

freeze: $-(50.0 \text{ g} \times 335 \text{ J/g}) = -16,750 \text{ J}$

$q = -4602 \text{ J} + (-16,750 \text{ J}) = -21,352 \text{ J} = \mathbf{-2.14 \times 10^4 \text{ J} \cdot \text{answer D}}$

A skipped the cooling leg: $-16,750 \text{ J} = -1.68 \times 10^4 \text{ J}$. B lost the sign: the water releases heat, so q is negative. C skipped the freezing plateau: $-4.60 \times 10^3 \text{ J}$. E cooled the liquid with ice's c: $50.0 \times 2.03 \times (-22.0) - 16,750 = -1.90 \times 10^4 \text{ J}$.

 **Sense check:** Freezing releases more than three times the heat of the cooling. ✓



Practice 4

total q = the sum of every segment crossed

given: 35.0 g ice at $-25.0\text{ }^{\circ}\text{C}$ $\rightarrow$ steam at $110.0\text{ }^{\circ}\text{C}$ $\cdot c(\text{ice}) 2.03 \cdot c(\text{water}) 4.184 \cdot c(\text{steam}) 1.9\text{ J/g}\cdot^{\circ}\text{C} \cdot$
 $\Delta H_{\text{fus}} 335 \cdot \Delta H_{\text{vap}} 2259\text{ J/g}$ $\cdot$ wanted: q in kJ

A **35.0 g** block of ice at **$-25.0\text{ }^{\circ}\text{C}$** is heated until it is steam at **$110.0\text{ }^{\circ}\text{C}$** . How much heat, in kJ, does it absorb?

- A. 96.2
- B. 115
- C. 108
- D. 1.08×10^5
- E. 28.1

✓ Practice 4 answer: C

given: 35.0 g ice, $-25.0\text{ }^{\circ}\text{C} \rightarrow$ steam, $110.0\text{ }^{\circ}\text{C}$ · c(ice) 2.03 · c(water) 4.184 · c(steam) 1.9 J/g· $^{\circ}\text{C}$ · ΔH_{fus} 335 · ΔH_{vap} 2259 J/g

warm ice: $35.0 \times 2.03 \times 25.0 = 1776\text{ J}$ · melt: $35.0 \times 335 = 11,725\text{ J}$

warm water: $35.0 \times 4.184 \times 100 = 14,644\text{ J}$ · boil: $35.0 \times 2259 = 79,065\text{ J}$

warm steam: $35.0 \times 1.9 \times 10.0 = 665\text{ J}$ · $q = 107,875\text{ J} = \mathbf{108\text{ kJ} \cdot \text{answer C}}$

A skipped the melt: $107,875 - 11,725 = 96.2\text{ kJ}$. B used $110.0\text{ }^{\circ}\text{C}$ as the steam's ΔT : $35.0 \times 1.9 \times 110.0 = 7315\text{ J}$, total 115 kJ. D left the total in joules: 1.08×10^5 . E stopped before boiling: $1776 + 11,725 + 14,644 = 28.1\text{ kJ}$.

 **Sense check:** Boiling alone is 79.1 kJ, nearly three-quarters of the total. ✓



Practice 5

heat absorbed = the sum of every segment crossed

given: 25.0 g ice at $-10.0\text{ }^{\circ}\text{C}$ · absorbs $1.500 \times 10^4\text{ J}$ · ends as liquid water · $c(\text{ice})\ 2.03$ · $c(\text{water})\ 4.184\text{ J/g}\cdot^{\circ}\text{C}$ · $\Delta H_{\text{fus}}\ 335\text{ J/g}$ · wanted: T_{final}

A **25.0 g** ice cube at **$-10.0\text{ }^{\circ}\text{C}$** absorbs **$1.500 \times 10^4\text{ J}$** and ends as liquid water. What is its final temperature, in $^{\circ}\text{C}$?

- A. 63.3
- B. 58.5
- C. 139
- D. 48.5
- E. 53.3

✓ Practice 5 answer: B

given: 25.0 g ice at $-10.0\text{ }^{\circ}\text{C}$ · absorbs 15,000 J · $c(\text{ice})\ 2.03$ · $c(\text{water})\ 4.184\text{ J/g}\cdot^{\circ}\text{C}$ · $\Delta H_{\text{fus}}\ 335\text{ J/g}$

warm ice: $25.0 \times 2.03 \times 10.0 = 507.5\text{ J}$ · melt: $25.0 \times 335 = 8375\text{ J}$ ·

left for the water: $15,000 - 507.5 - 8375 = 6117.5\text{ J}$

$$\Delta T = \frac{6117.5\text{ J}}{25.0\text{ g} \times 4.184\text{ J/(g}\cdot^{\circ}\text{C)}} = 58.5\text{ }^{\circ}\text{C} \cdot T_{\text{final}} = 0\text{ }^{\circ}\text{C} + 58.5\text{ }^{\circ}\text{C} = \mathbf{58.5\text{ }^{\circ}\text{C} \cdot \text{answer B}}$$

A skipped the ice leg: $6625 \div 104.6 = 63.3$. C skipped the melt: $14,492.5 \div 104.6 = 139$, past the boiling point. D started the water at $-10.0\text{ }^{\circ}\text{C}$: $-10.0 + 58.5 = 48.5$. E warmed the ice with water's c : $5579 \div 104.6 = 53.3$.

 **Sense check:** Melting takes over half the heat. The water warms from $0\text{ }^{\circ}\text{C}$, not from the ice's start. ✓



Check yourself

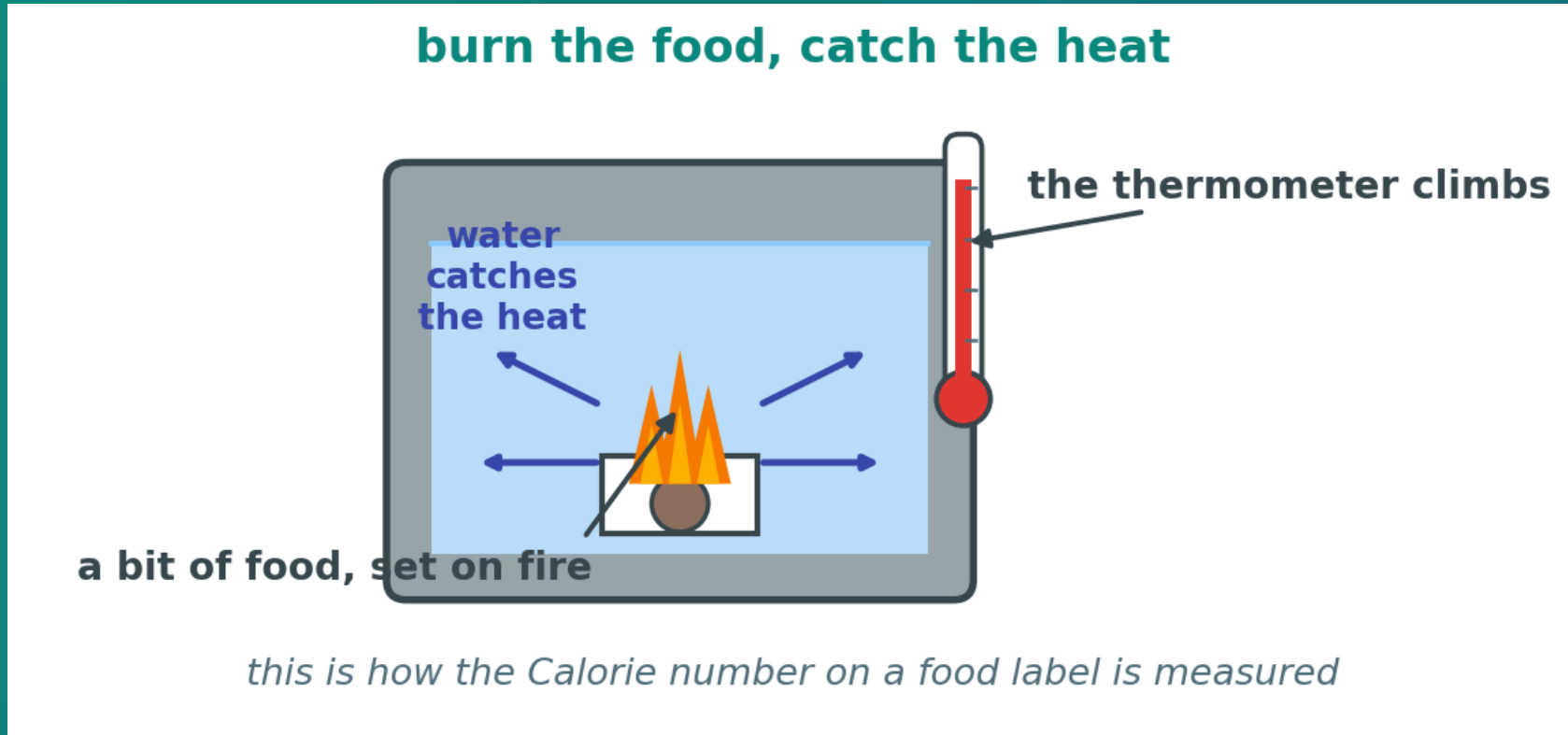
1. A path crosses a flat plateau, then a rising slope. Which equation does each part use?
Why does the plateau carry no ΔT ?
2. Water at 100 °C and steam at 100 °C are the same temperature. Why does boiling the water still take heat?

Drop a hot metal into cool water and heat leaves the metal and enters the water until both settle at one shared temperature. Mixing two bodies and reading the thermometer is **calorimetry**: the heat lost equals the heat gained.

8 • Calorimetry

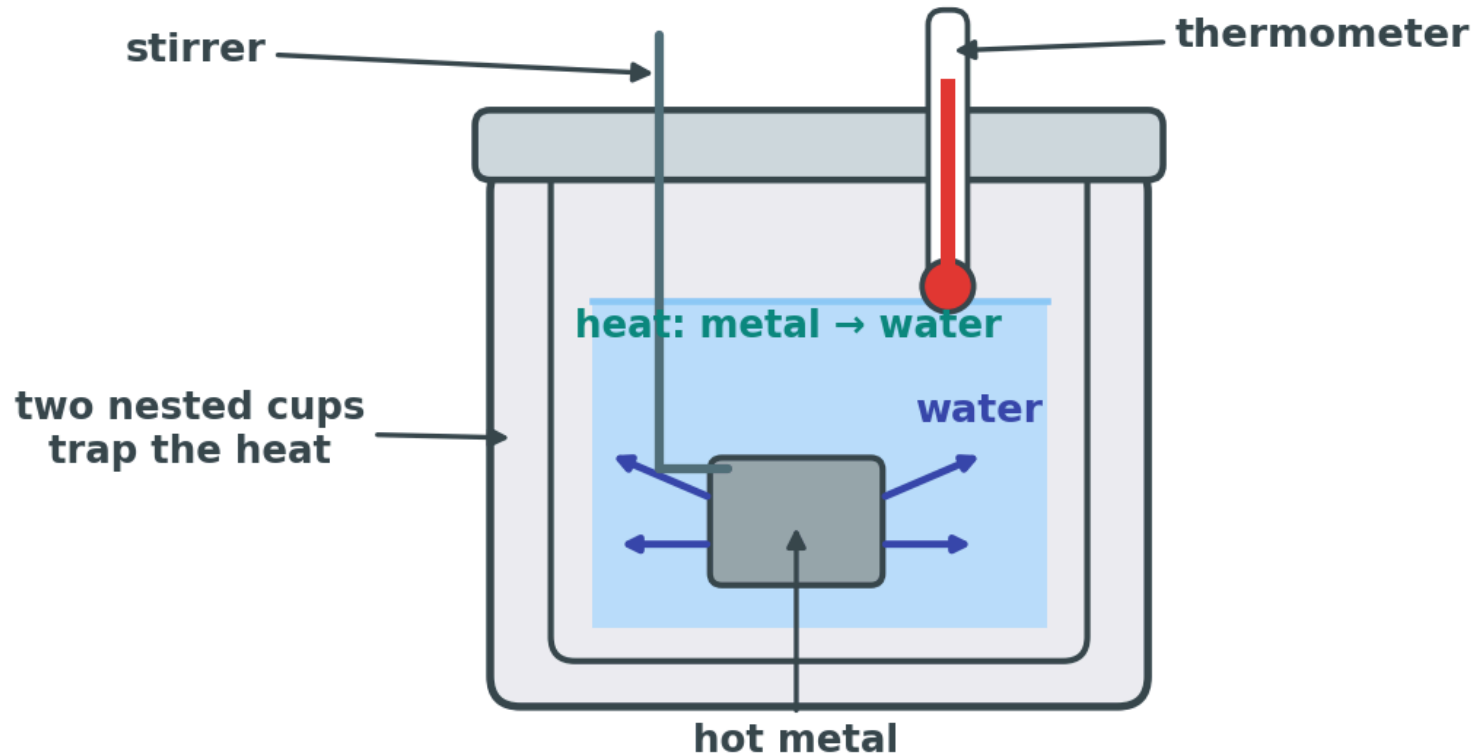
Use an insulated cup to make energy conservation visible (the heat one body loses equals the heat another gains) and solve for a final temperature or a reaction's enthalpy, with the answer's final temperature landing between the two starting temperatures.

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.

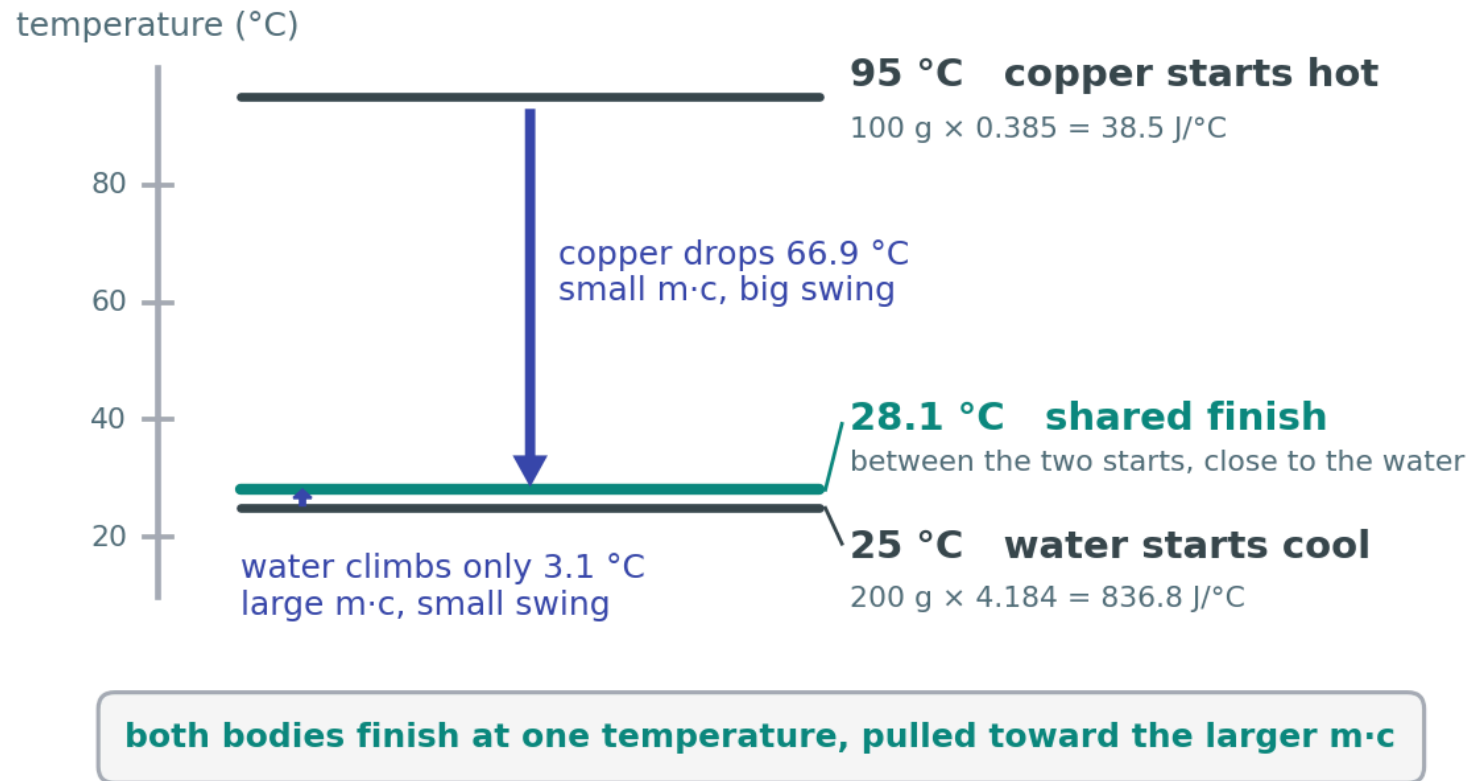
Energy is conserved in the cup

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



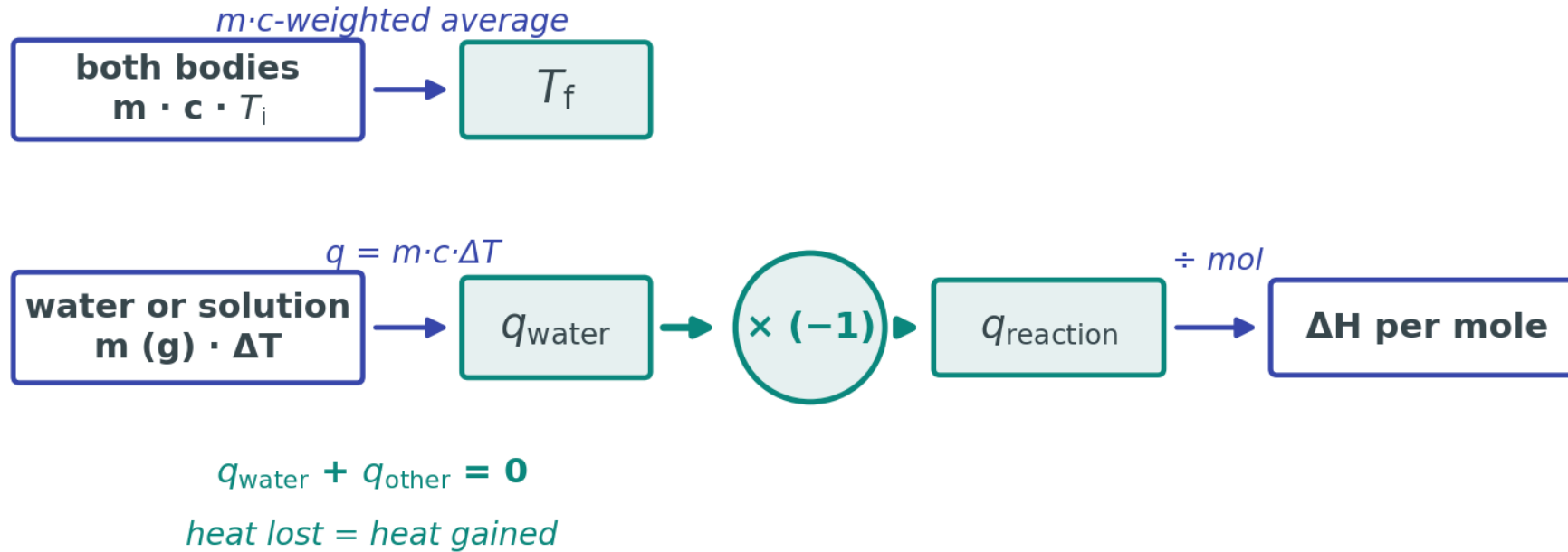
Both bodies end at the same temperature. It lands between the two starting temperatures, pulled toward whichever body carries the larger m·c: usually the water.



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.

Two routes on the map



Top row: with T_f unknown, the $m \cdot c$ -weighted average gives it in one move. Bottom row: the solution's $q = m \cdot c \cdot \Delta T$, one sign flip for the reaction's heat, then $\div$ moles for ΔH .



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

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

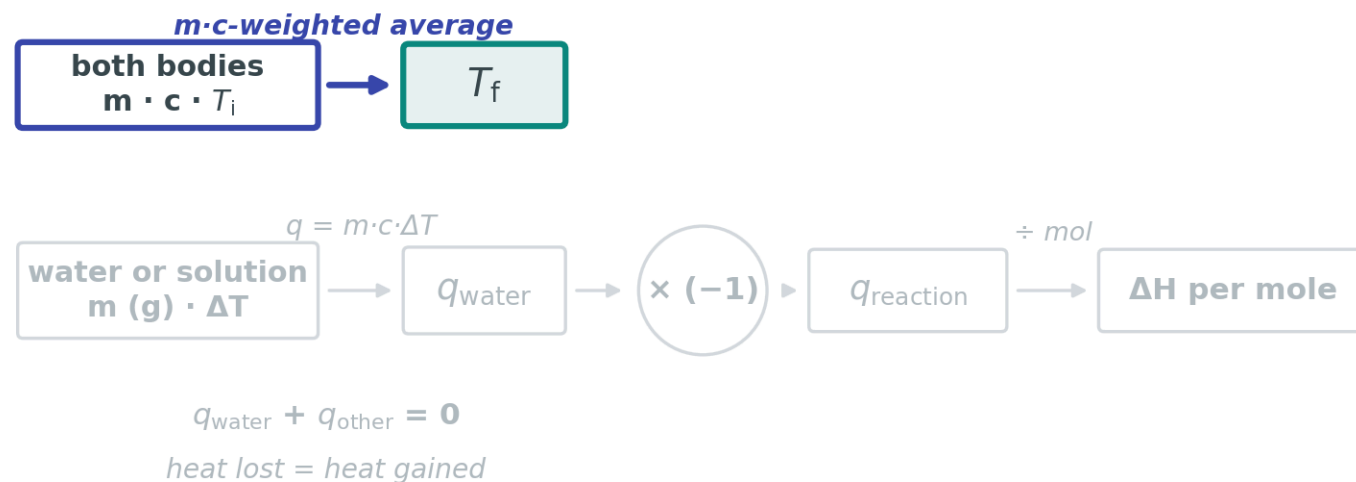
Step 3 · Check

🧠 **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: $T_f = 50.0\text{ °C}$



🧠 **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

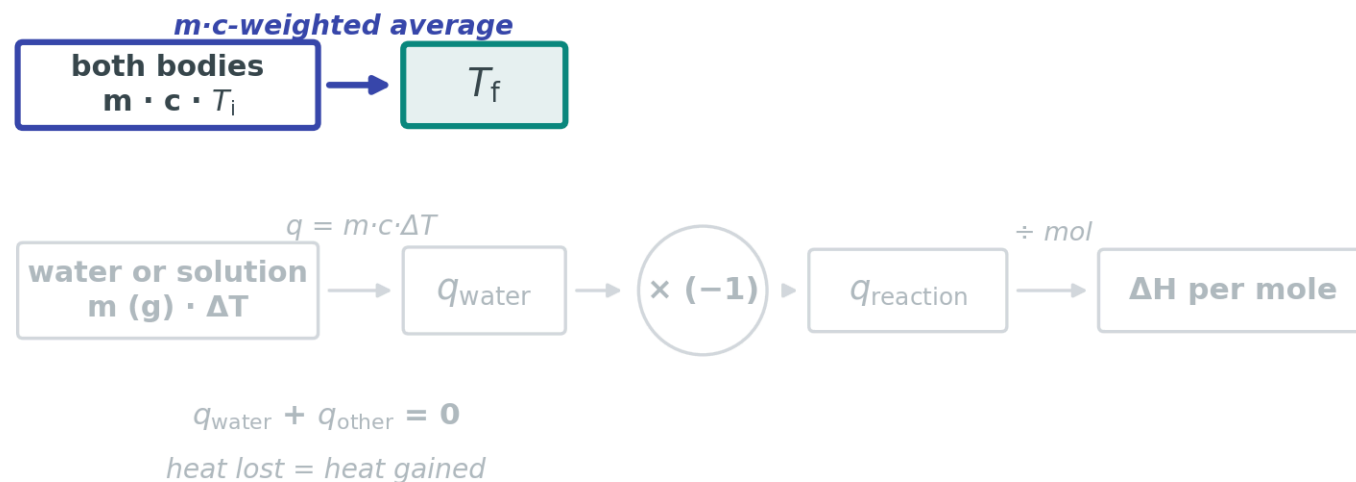
$$T_f = \frac{44.9 \times 90 + 627.6 \times 20}{44.9 + 627.6} = 24.7 \text{ °C}$$

🧠 **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 $m \cdot c$ 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.

✓



Worked example 3: enthalpy from a temperature rise

$$q = m \cdot c \cdot \Delta T, \text{ then } \Delta H = -q / \text{moles}$$

given: 100.0 g solution · 20.0 → 26.8 °C · 0.0500 mol reacted · wanted: ΔH per mole

An acid and a base neutralize in a cup holding **100.0 g** of solution. The temperature climbs from 20.0 to 26.8 °C, and **0.0500 mol** of water forms. Find ΔH per mole.

✓ Worked example 3: solution

$$q = m \cdot c \cdot \Delta T, \text{ then } \Delta H = -q / \text{moles}$$

$$100.0 \text{ g solution} \cdot \Delta T = +6.8 \text{ }^{\circ}\text{C} \cdot 0.0500 \text{ mol}$$

Two moves are needed.

Step 1 • Heat gained by the solution

$$q = 100.0 \text{ g} \times 4.184 \text{ J/g}\cdot^{\circ}\text{C} \times 6.8 \text{ }^{\circ}\text{C} = \mathbf{2845 \text{ J}}$$

Step 2 • Heat per mole of reaction

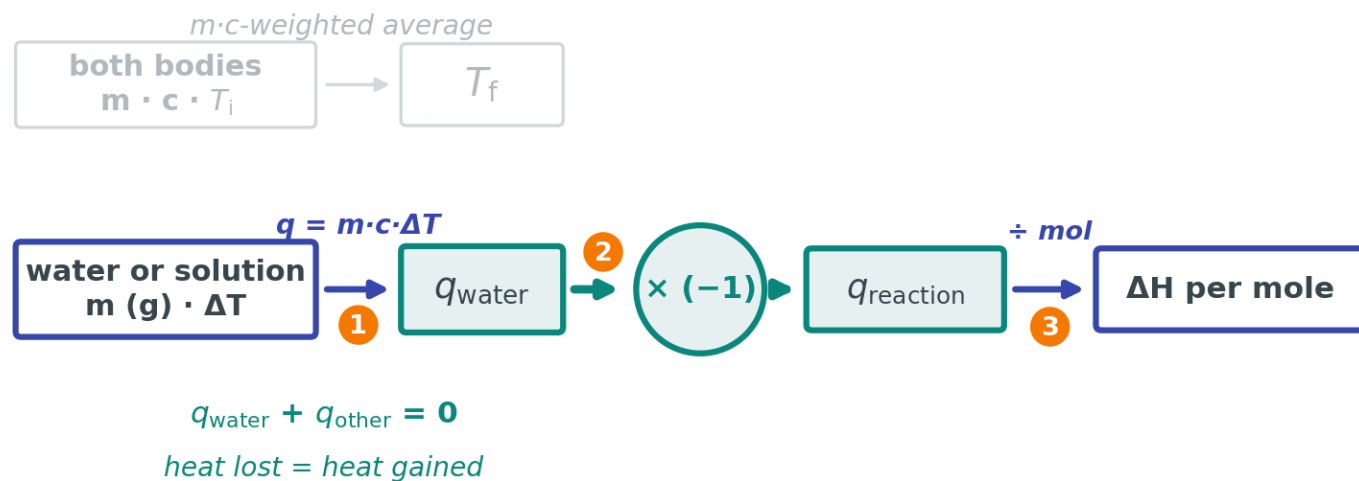
$$\Delta H = \frac{-2845 \text{ J}}{0.0500 \text{ mol}} = -56,900 \text{ J/mol} = \mathbf{-56.9 \text{ kJ/mol}}$$

 **Sense check:** The solution warmed, so the reaction released heat: ΔH is negative, exothermic. ✓

✓ Worked example 3: the route on the map

$$q_{\text{reaction}} = -q_{\text{solution}} \cdot \Delta H = q_{\text{reaction}} \div \text{moles}$$

given: 100.0 g solution, 20.0 → 26.8 °C · 0.0500 mol · found: $\Delta H = -56.9 \text{ kJ/mol}$



 **Sense check:** Step 1 of the solution is arrow 1: the solution's q . Step 2 is arrows 2 and 3: flip the sign, then divide by the moles. ✓

Two calorimeters, two jobs

bomb calorimeter: sealed steel, constant volume

combustion and food Calories · the whole rig warms: $q = C_{\text{cal}} \cdot \Delta T$

coffee-cup calorimeter: open cup, constant pressure

neutralization, dissolving, mixing · the heat measured here equals ΔH · a dissolving salt's heat, cold pack or hot pack, gets its number this way

Combustion runs in the sealed bomb at constant volume. Solution chemistry runs in the open cup at constant pressure, where q equals ΔH . Burn it in the bomb; mix it in the cup.



Worked example 4: a bomb calorimeter run

$$q_{\text{released}} = C_{\text{cal}} \cdot \Delta T$$

given: 1.0 g fructose burned · $C_{\text{cal}} = 9.90 \text{ kJ/}^\circ\text{C}$ · $\Delta T = +1.58 \text{ }^\circ\text{C}$ · wanted: heat released

1.0 g of fructose burns in a bomb calorimeter whose heat capacity is $C_{\text{cal}} = \mathbf{9.90 \text{ kJ/}^\circ\text{C}}$. The temperature climbs **1.58 °C**. How much heat did the combustion release?

✓ Worked example 4: solution

$$q_{\text{released}} = C_{\text{cal}} \cdot \Delta T$$

$$1.0 \text{ g fructose} \cdot C_{\text{cal}} = 9.90 \text{ kJ/}^{\circ}\text{C} \cdot \Delta T = +1.58 \text{ }^{\circ}\text{C}$$

One heat capacity for the whole rig

Multiply by the temperature rise

$$q = \frac{9.90 \text{ kJ}}{1 \text{ }^{\circ}\text{C}} \times 1.58 \text{ }^{\circ}\text{C} = 15.6 \text{ kJ}$$

Set the direction

 **Sense check:** The calorimeter warmed, so the combustion released the heat: $q_{\text{rxn}} = -15.6 \text{ kJ}$ for one gram of fructose. That is $15.6 \div 4.184 = 3.74$, nearly 4 food Calories per gram: the carbohydrate number on every label, measured. ✓

⚠ 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

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 1 · 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 2

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

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 2 · 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 3

$q_{\text{released}} = n \times |\Delta H|$, then the calorimeter constant $C_{\text{cal}} = q / \Delta T$

given: 0.562 g carbon burns · $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$ · $\Delta H = -393.5 \text{ kJ/mol}$ · $26.74^\circ\text{C} \rightarrow 27.93^\circ\text{C}$ · wanted: C_{cal}

0.562 g of carbon burns in a bomb calorimeter: $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$, $\Delta H = -393.5 \text{ kJ/mol}$. The temperature climbs from 26.74°C to 27.93°C . What is the calorimeter constant, in kJ per $^\circ\text{C}$?

- A. 15.5
- B. 18.4
- C. 186
- D. 0.659

✓ Practice 3 · answer: A



$$0.562 \text{ g C burned} \cdot \Delta T = 27.93 - 26.74 = +1.19 \text{ }^\circ\text{C}$$

$$q = 0.562 \text{ g} \times \frac{1 \text{ mol}}{12.01 \text{ g}} \times \frac{393.5 \text{ kJ}}{1 \text{ mol}} = 18.4 \text{ kJ} \rightarrow C_{\text{cal}} = \frac{18.4 \text{ kJ}}{1.19 \text{ }^\circ\text{C}}$$

= 15.5 kJ/°C (answer A)

B stopped at the heat: 18.4 kJ is q, not a calorimeter constant. C skipped the mole conversion: $0.562 \times 393.5 \div 1.19 = 186$ treats the grams as moles. D divided by the final temperature: $18.4 \div 27.93 = 0.659$; the divisor is $\Delta T = 1.19 \text{ }^\circ\text{C}$.

 **Sense check:** A reaction with a known ΔH calibrates the rig: 18.4 kJ went in and the temperature rose $1.19 \text{ }^\circ\text{C}$, so the calorimeter soaks up 15.5 kJ per degree. That one number is the calorimeter constant; run the same calibration on a coffee cup and it goes by the cup constant, K_{cup} . ✓



Check yourself

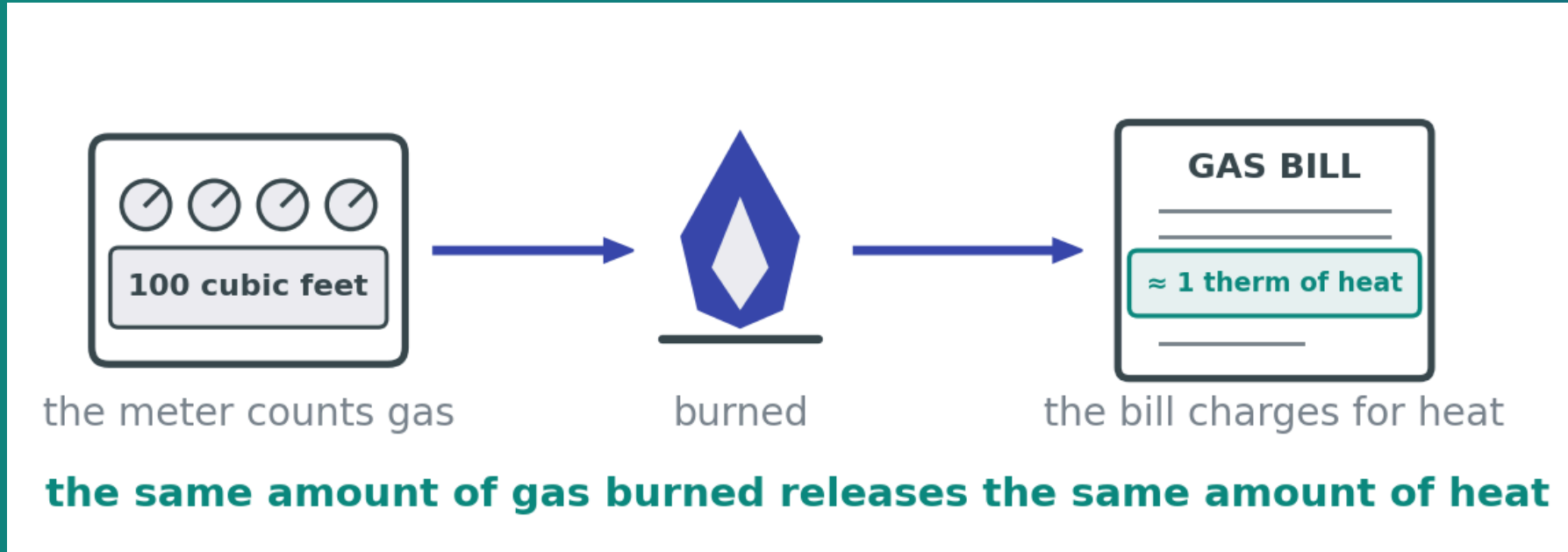
1. A hot metal is dropped into cool water in an insulated cup. Write the conservation law that relates the heat the metal loses to the heat the water gains.
2. A final temperature comes out below both starting temperatures. What went wrong, and what does the correct answer always sit between?

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.

9 • 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.

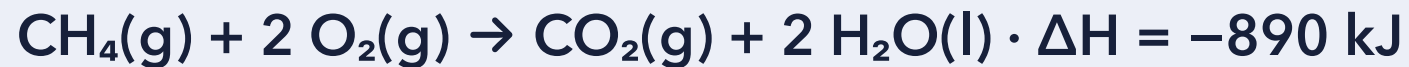
The gas bill charges for heat



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)

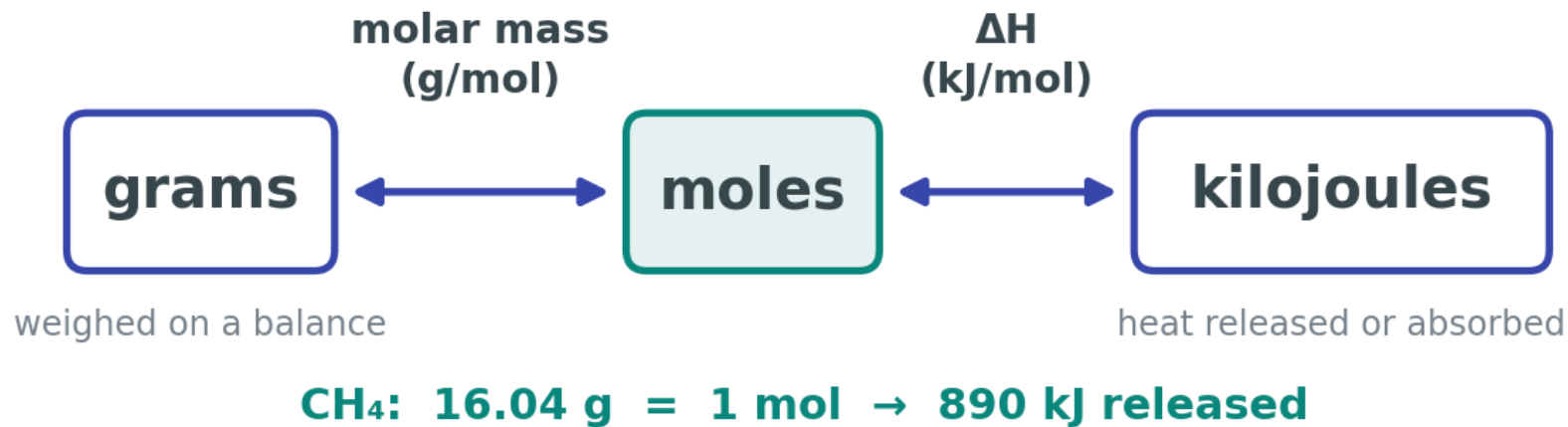
Scale the coefficients, scale ΔH

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.



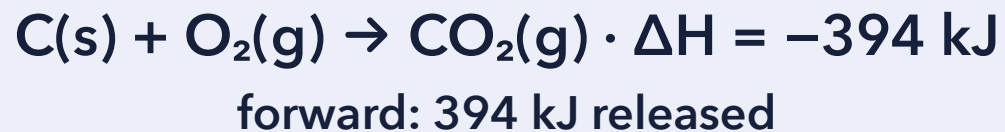
Worked example 1: reversing an equation



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

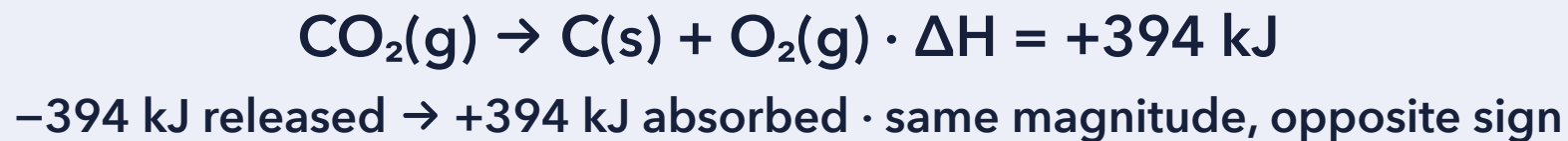
Carbon burns to carbon dioxide, releasing 394 kJ. What is ΔH for the reverse, $\text{CO}_2\text{(g)} \rightarrow \text{C(s)} + \text{O}_2\text{(g)}$?

✓ 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).

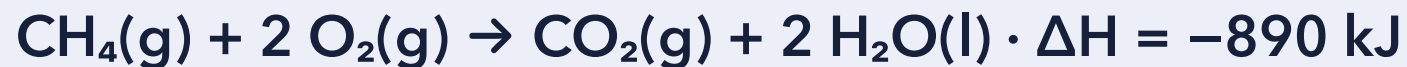
$$\Delta H = \underline{\hspace{1cm}} (\underline{\hspace{1cm}} \times -198.2 \text{ kJ}) = \underline{\hspace{1cm}} \text{ kJ}$$

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

 **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. ✓



Worked example 3: grams of fuel to kilojoules

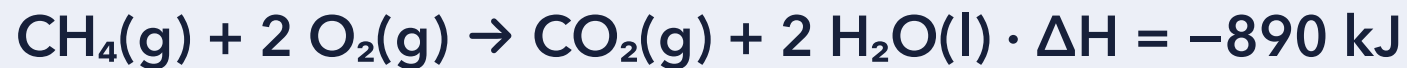


given: 40.0 g CH₄ burned · molar mass CH₄ = 16.04 g/mol · wanted: ΔH

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

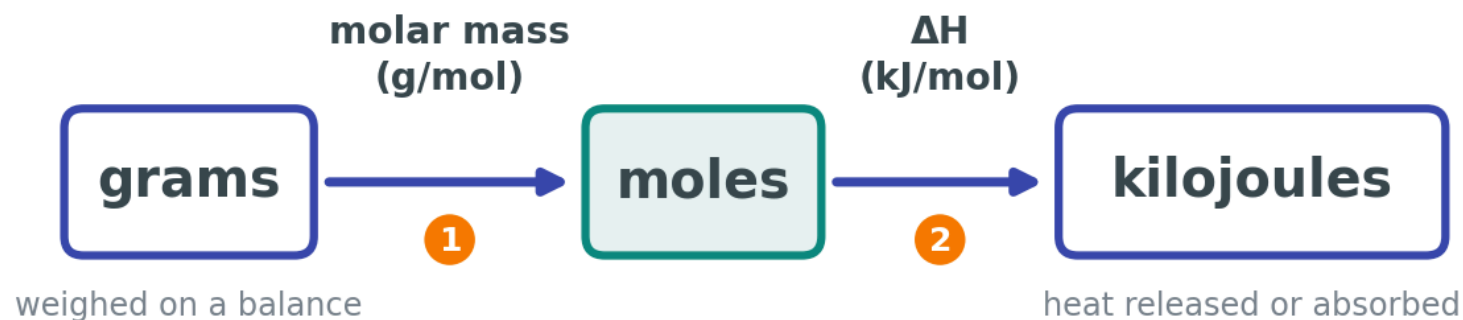
$$40.0 \text{ g} \times \frac{1 \text{ mol}}{16.04 \text{ g}} \times \frac{-890 \text{ kJ}}{1 \text{ mol}} = -2220 \text{ kJ}$$

🧠 **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 $\text{SO}_2 = 64.07 \text{ g/mol}$ · wanted: heat evolved in kJ

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

Same chain as the methane problem, with one difference: -198.2 kJ does not belong to 1 mol of SO_2 . Tempting: use it per mole anyway. Test it.

✓ Worked example 4: solution



$$87.9 \text{ g SO}_2 \cdot 64.07 \text{ g/mol} \cdot -198.2 \text{ kJ per 2 mol SO}_2$$

Step 1 • Write ΔH per mole of SO_2

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

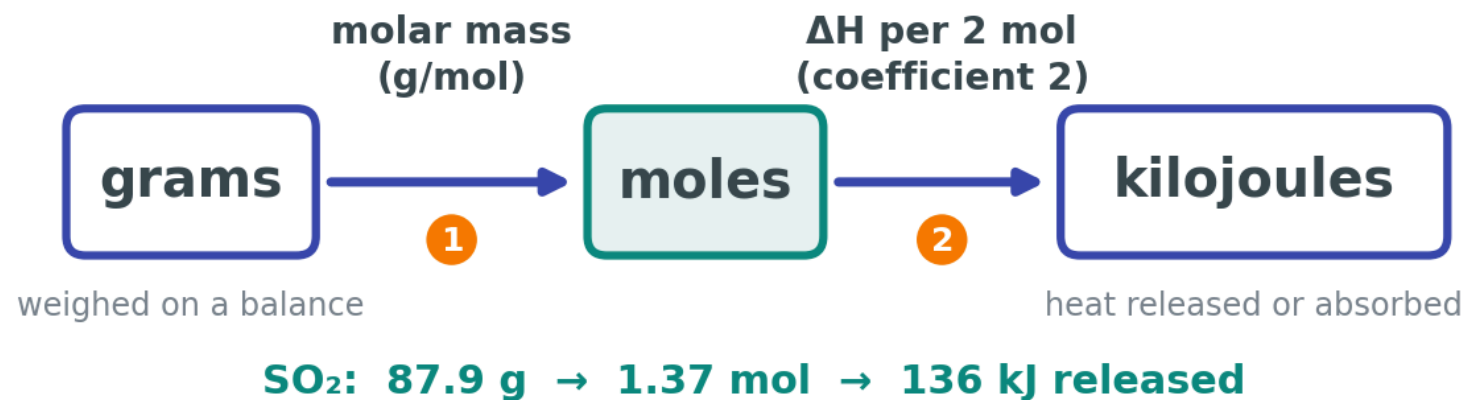
$$87.9 \text{ g} \times \frac{1 \text{ mol SO}_2}{64.07 \text{ g}} \times \frac{-198.2 \text{ kJ}}{2 \text{ mol SO}_2} = -136 \text{ kJ}$$

🧠 **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 kJ per 2 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 1

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 1 · 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 2

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 2 · 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 3

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 3 · 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

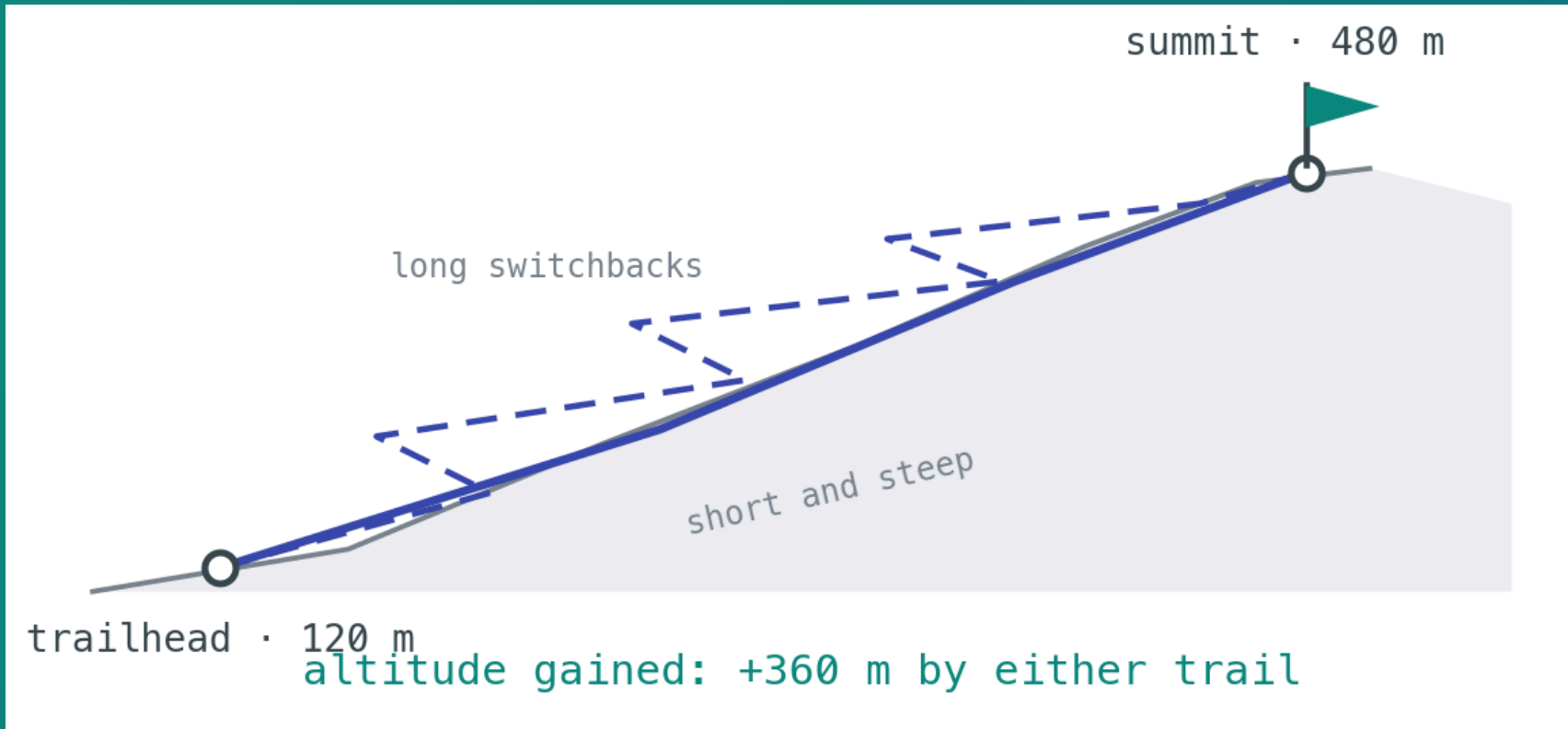
1. An equation is reversed and its coefficients tripled. What two changes apply to ΔH ?
2. A reaction releases 500 kJ per mole. Write the conversion factor that turns moles of it into kilojoules, and state which unit cancels.

A single reaction's ΔH can be reversed and scaled. Adding several such equations, each reversed or scaled to line up, builds the ΔH of a reaction never measured directly: **Hess's law**.

10 • Hess's Law

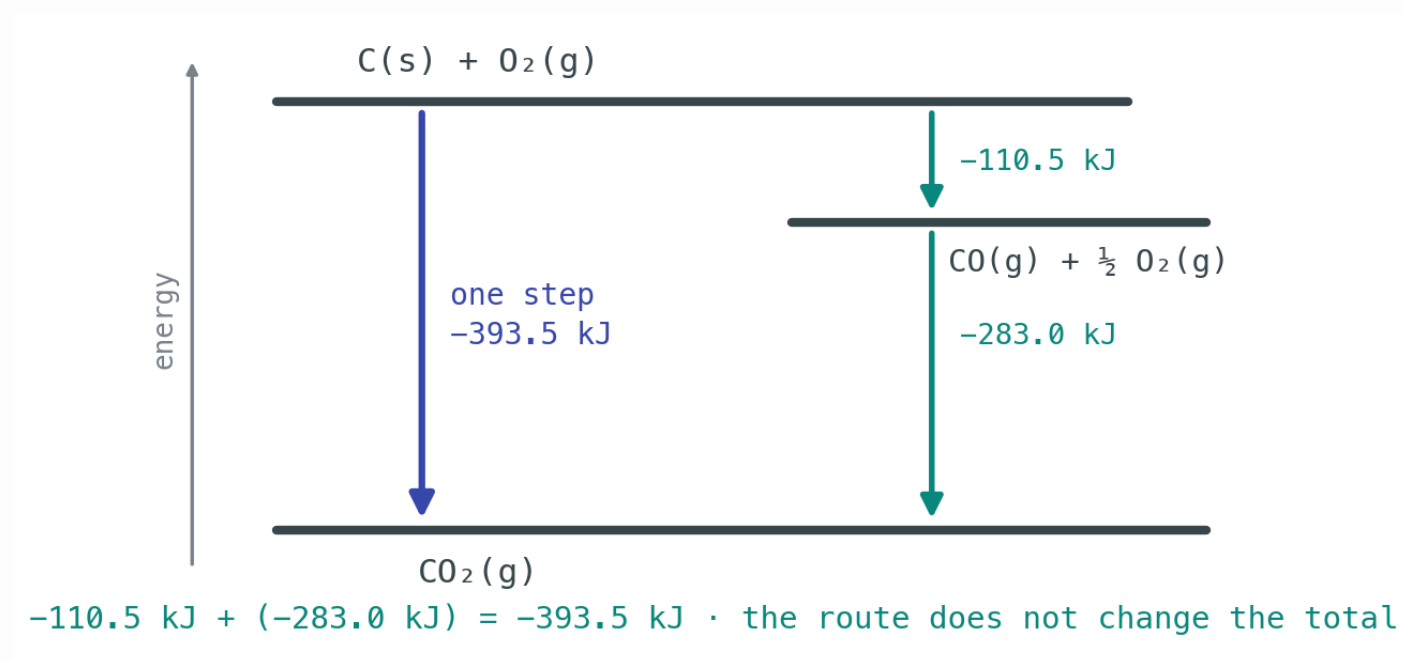
Use Hess's law to find the enthalpy of an overall reaction, rearranging the given thermochemical equations, reversing and scaling each ΔH in step, then adding so that everything cancels except the target.

Two trails, one altitude gain



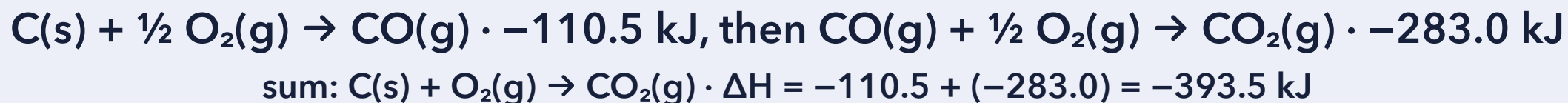
Two trails climb from the same trailhead to the same summit. The altitude gained is identical for both. Enthalpy changes work the same way.

Enthalpy is a state function



A state function depends only on the start and the end, never on the route. Burning carbon to CO_2 releases 393.5 kJ whether it takes one step or two.

Hess's law: the steps' ΔH values add



When equations add up to a target reaction, their ΔH values add up to its ΔH . This gives ΔH for reactions no calorimeter can isolate.

Rearrange each given before adding

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

reversed: $\Delta H \times (-1)$ · scaled by n: $\Delta H \times n$ · a reversed and doubled equation needs both moves

Given equations rarely line up as printed. Reverse or scale each one until its species sit where the target needs them, and change its ΔH in step. Say the rule as you work.



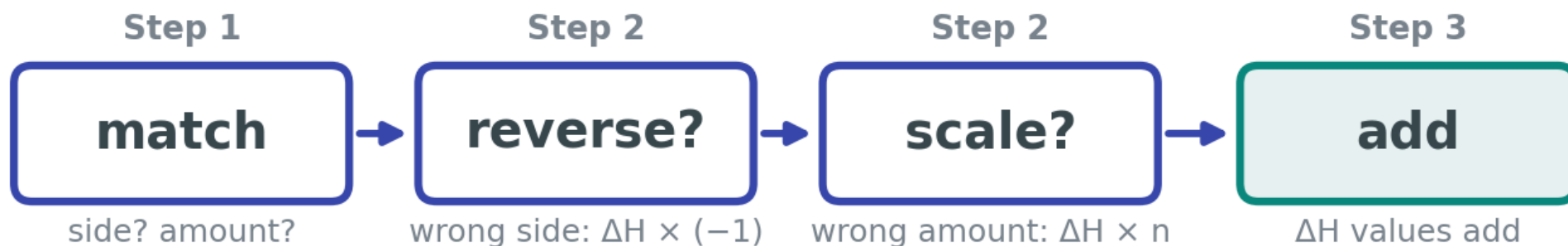
The method

$\text{C(s)} + \frac{1}{2} \text{O}_2\text{(g)} \rightarrow \text{CO(g)}$: burning carbon always makes CO_2 too · $2 \text{C(s)} + \text{H}_2\text{(g)} \rightarrow \text{C}_2\text{H}_2\text{(g)}$: will not run at all

the indirect route: no calorimeter sees these, so their ΔH is built from combustions that measure cleanly

1. **Match each given to the target.** Note the side and coefficient each needs.
2. **Reverse and scale.** Flip the sign for a reversal; multiply ΔH with the coefficients.
3. **Add and cancel.** Sum equations and ΔH values; shared species cancel.

The method on the map

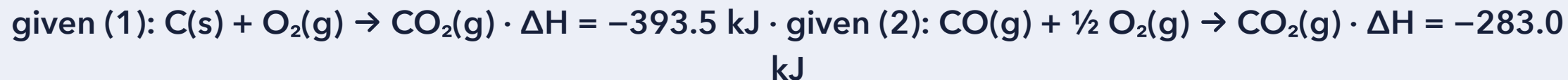


Every given runs the route. Only the target survives the sum.

Thermochemical equations already reversed and scaled one equation at a time. Hess's law asks both questions of every given, then adds the adjusted equations and their ΔH values.



Worked example 1: combining two equations



Burning carbon always makes some CO_2 alongside the CO , so ΔH for this reaction cannot be measured directly. Find it from the two given equations.

✓ Worked example 1: solution

target: $\text{C(s)} + \frac{1}{2} \text{O}_2\text{(g)} \rightarrow \text{CO(g)}$

(1) $\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} \cdot -393.5 \text{ kJ}$ · (2) $\text{CO(g)} + \frac{1}{2} \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} \cdot -283.0 \text{ kJ}$

Step 1 · Match each given to the target

Step 2 · Reverse and scale

reversed (2): $\text{CO}_2\text{(g)} \rightarrow \text{CO(g)} + \frac{1}{2} \text{O}_2\text{(g)} \cdot \Delta\text{H} = +283.0 \text{ kJ}$

sign flips: -283.0 becomes $+283.0$ · coefficients unchanged, so no scaling

Step 3 · Add and cancel

$$\Delta\text{H} = -393.5 \text{ kJ} + 283.0 \text{ kJ} = -110.5 \text{ kJ}$$

🧠 **Sense check:** CO_2 and $\frac{1}{2} \text{O}_2$ appear on both sides of the sum and cancel, leaving exactly the target. A -110.5 step then a -283.0 step totals -393.5 , the one-step value. ✓

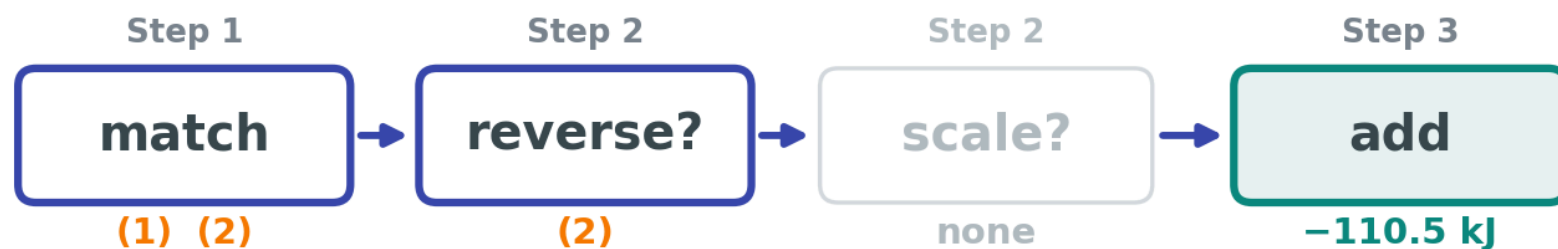
✓ Worked example 1: the route on the map



Match on the map

Reverse and scale on the map

Only (2) is reversed: -283.0 kJ becomes $+283.0 \text{ kJ}$. Every coefficient already matches, so nothing is scaled.



🧠 **Sense check:** Step 3 adds -393.5 kJ and $+283.0 \text{ kJ}$: -110.5 kJ . Scale stays grey, since no given needed it. ✓



Worked example 2: reversing and scaling a given



Both givens are written per mole of carbon, and the target contains no carbon at all. Find ΔH for the target.

✓ Worked example 2: solution



Step 1 · Match each given to the target

Step 2 · Reverse and scale



Step 3 · Add and cancel

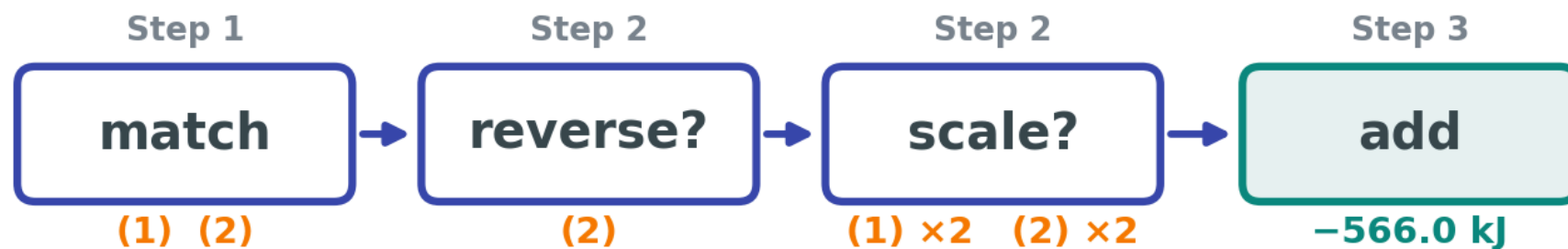
$$\Delta H = 221.0 \text{ kJ} + (-787.0 \text{ kJ}) = -566.0 \text{ kJ}$$

 **Sense check:** 2 C and one O₂ cancel in the sum, leaving the target. Burning CO to CO₂ releases 283.0 kJ per mole, and 2 × 283.0 = 566.0. ✓

✓ Worked example 2: the route on the map

target: $2 \text{ CO(g)} + \text{O}_2\text{(g)} \rightarrow 2 \text{ CO}_2\text{(g)}$

(1) $\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} \cdot -393.5 \text{ kJ}$ · (2) $\text{C(s)} + \frac{1}{2} \text{ O}_2\text{(g)} \rightarrow \text{CO(g)} \cdot -110.5 \text{ kJ}$



🧠 **Sense check:** (2) takes both moves: CO sits on the wrong side, and the target needs 2 of it. (1) skips the reversal but still doubles. Both givens light up the scale box. ✓

Your turn: sulfur dioxide to sulfur trioxide

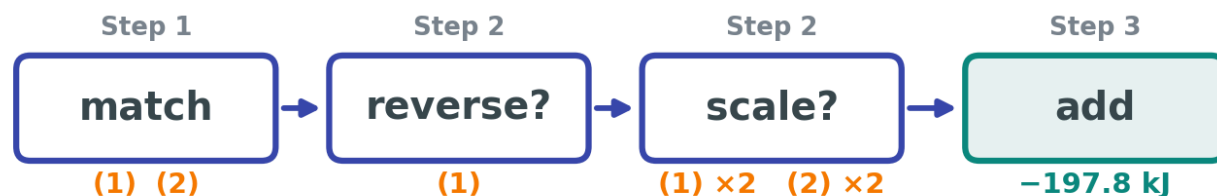


Reverse and double (1), double (2), then add.

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

$$\Delta\text{H} = -(2 \times (-296.8 \text{ kJ})) + 2 \times (-395.7 \text{ kJ}) = 593.6 \text{ kJ} - 791.4 \text{ kJ} = \mathbf{-197.8 \text{ kJ}}$$

 **Sense check:** Both routes start from sulfur: forming 2 SO₃ directly releases 791.4 kJ, forming 2 SO₂ releases 593.6 kJ, and the target is the step between them. ✓



Where this goes wrong

target: $\text{C(s)} + \frac{1}{2} \text{O}_2\text{(g)} \rightarrow \text{CO(g)}$ · correct: reverse (2), add · $\Delta\text{H} = -110.5 \text{ kJ}$

(1) $\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$ · -393.5 kJ · (2) $\text{CO(g)} + \frac{1}{2} \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$ · -283.0 kJ

Reversing without flipping the sign. Adding -393.5 and -283.0 unchanged gives -676.5 kJ , the heat of two combustions, not the target. Reversed, (2) contributes $+283.0 \text{ kJ}$.

Scaling the equation but not ΔH . A given doubled to fit the target doubles its ΔH too. A reversed and doubled given needs both the flip and the $\times 2$.

Adding before the species line up. If anything besides the target species survives the sum, a given still needs reversing or scaling. Fix the equations first; the ΔH values follow.



Practice 1



What is ΔH , in kJ, for the target reaction?

- A. +294.6
- B. +66.4
- C. +180.5
- D. -294.6

✓ Practice 1 · answer: B



(1) $+180.5 \text{ kJ}$ · (2) -114.1 kJ · 2 NO cancels when the givens are added as written

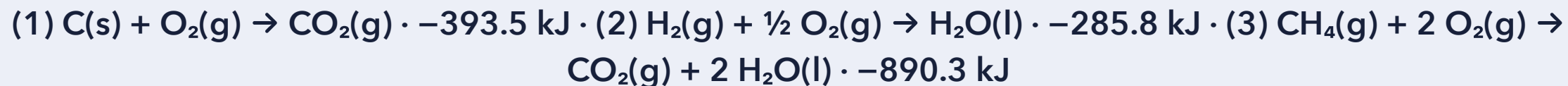
$$\Delta H = 180.5 \text{ kJ} + (-114.1 \text{ kJ}) = \mathbf{+66.4 \text{ kJ (answer B)}}$$

Neither given needs reversing or scaling. A dropped the minus on (2): $180.5 + 114.1 = +294.6$. C stopped at the first given: $+180.5 \text{ kJ}$ never adds step (2). D flipped (1) instead: $-180.5 - 114.1 = -294.6$.

🧠 **Sense check:** Forming NO absorbs more heat than the NO-to-NO₂ step releases, so the overall ΔH stays positive.
✓

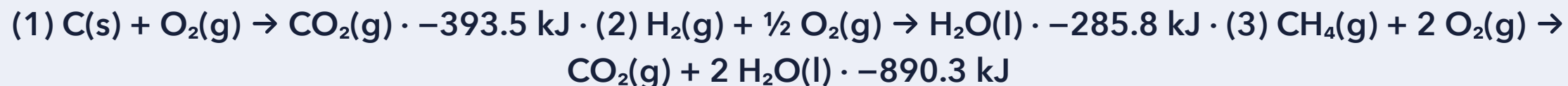


Worked example 3: three given equations



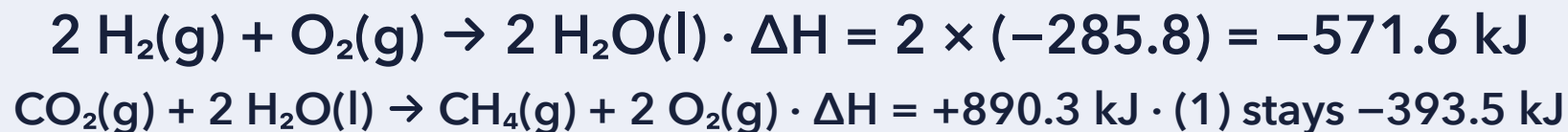
Carbon and hydrogen do not combine cleanly to methane in any calorimeter, but all three combustions are easy to measure. Find ΔH for the target.

✓ Worked example 3: rearranging the givens



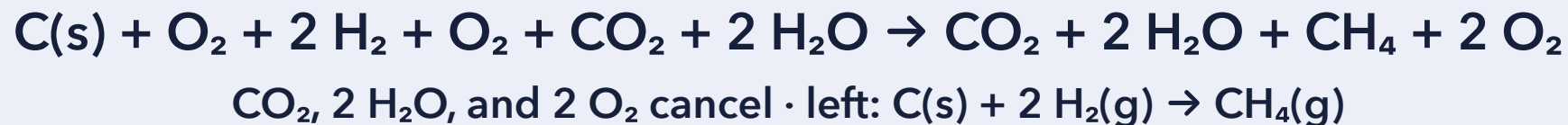
Step 1 · Match each given to the target

Step 2 · Reverse and scale



 **Sense check:** Each ΔH moved with its equation: $\times 2$ for the doubling, a sign flip for the reversal, no change for the equation kept as written. ✓

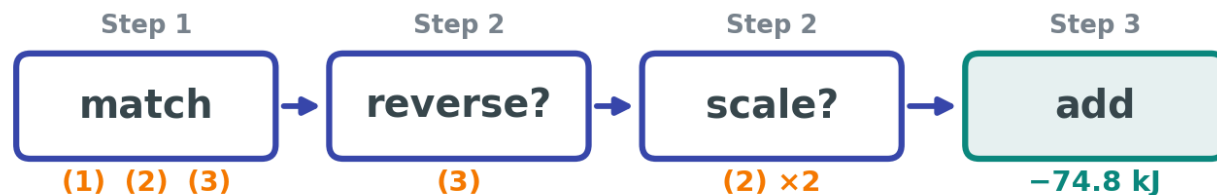
✓ Worked example 3: adding the three equations



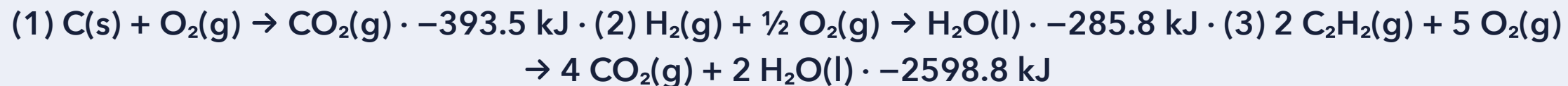
Step 3 · Add and cancel

$$\Delta H = -393.5 \text{ kJ} + (-571.6 \text{ kJ}) + 890.3 \text{ kJ} = \mathbf{-74.8 \text{ kJ}}$$

 **Sense check:** Burning the elements releases 965.1 kJ; unburning methane costs 890.3 kJ. The difference, 74.8 kJ, is released. ✓



Worked example 4: halving a given equation



Carbon and hydrogen will not combine directly to acetylene, so its formation enthalpy comes from three combustions. Read (3) closely: its -2598.8 kJ belongs to **two** moles of acetylene, and the target makes one.

✓ Worked example 4: solution

target: $2 \text{ C(s, graphite)} + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_2(\text{g})$

(1) $\text{C} + \text{O}_2 \rightarrow \text{CO}_2 \cdot -393.5 \text{ kJ}$ · (2) $\text{H}_2 + \frac{1}{2} \text{ O}_2 \rightarrow \text{H}_2\text{O(l)} \cdot -285.8 \text{ kJ}$ · (3) $2 \text{ C}_2\text{H}_2 + 5 \text{ O}_2 \rightarrow 4 \text{ CO}_2 + 2 \text{ H}_2\text{O(l)}$
· -2598.8 kJ

Step 1 · Match each given to the target

Step 2 · Reverse and scale

$2 \text{ CO}_2(\text{g}) + \text{H}_2\text{O(l)} \rightarrow \text{C}_2\text{H}_2(\text{g}) + \frac{5}{2} \text{ O}_2(\text{g}) \cdot \Delta\text{H} = +2598.8/2 = +1299.4 \text{ kJ}$

reversed: the sign flips · halved: ΔH halves · flipping without halving carries $+2598.8 \text{ kJ}$ and lands at $+1526.0 \text{ kJ}$, the answer for two moles

Step 3 · Add and cancel

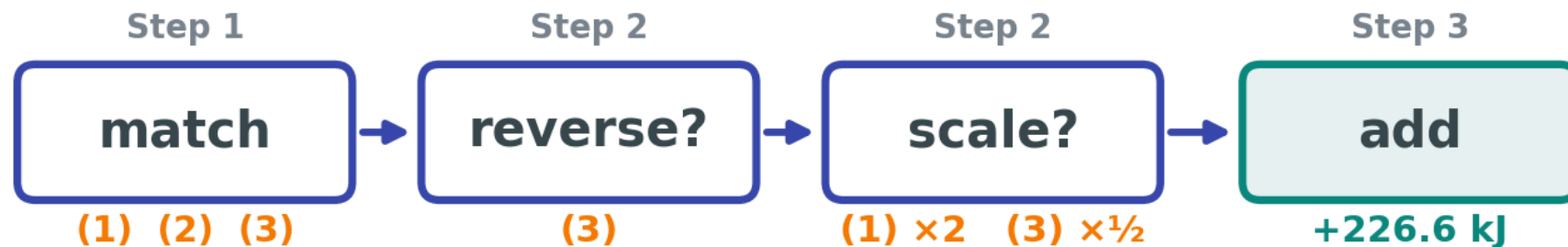
$$\Delta\text{H} = 2 \times (-393.5) + (-285.8) + 1299.4 = \mathbf{+226.6 \text{ kJ}}$$

 **Sense check:** 2 CO_2 , H_2O , and $\frac{5}{2} \text{ O}_2$ cancel, leaving exactly the target. Positive: acetylene sits 226.6 kJ above its elements, and a welding torch gets that stored energy back. ✓

✓ Worked example 4: the route on the map

target: $2 \text{ C(s, graphite)} + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_2(\text{g})$

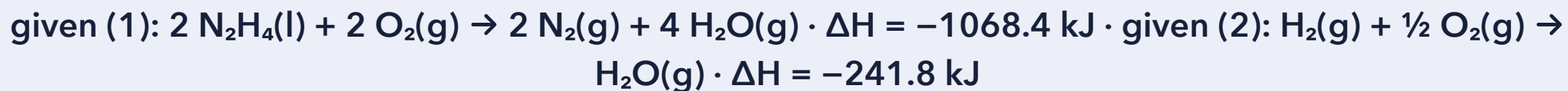
(1) $\text{C} + \text{O}_2 \rightarrow \text{CO}_2 \cdot -393.5 \text{ kJ}$ · (2) $\text{H}_2 + \frac{1}{2} \text{O}_2 \rightarrow \text{H}_2\text{O(l)} \cdot -285.8 \text{ kJ}$ · (3) $2 \text{ C}_2\text{H}_2 + 5 \text{ O}_2 \rightarrow 4 \text{ CO}_2 + 2 \text{ H}_2\text{O(l)}$
· -2598.8 kJ



🧠 **Sense check:** A scale factor can be less than 1. (3) is written for 2 mol acetylene, so it reverses and halves; (1) doubles; (2) passes straight to the sum. ✓



Practice 2



Hydrazine, a rocket fuel, cannot be made directly from its elements. What is ΔH , in kJ, for the target reaction?

- A. +50.6
- B. +534.2
- C. +584.8
- D. -1017.8
- E. +292.4

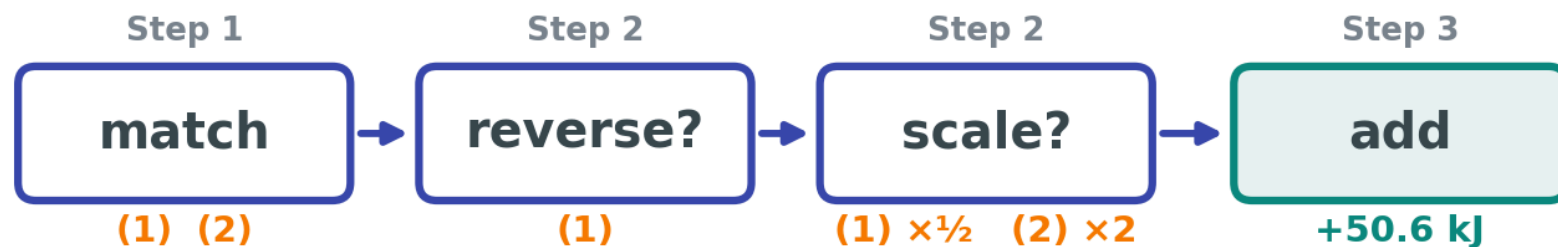
✓ Practice 2 · answer: A

reverse, halve (1): $\text{N}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g}) \rightarrow \text{N}_2\text{H}_4(\text{l}) + \text{O}_2(\text{g}) \cdot \Delta\text{H} = 1068.4/2 = +534.2 \text{ kJ}$

double (2): $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{g}) \cdot \Delta\text{H} = 2 \times (-241.8) = -483.6 \text{ kJ}$ · O_2 and $2 \text{H}_2\text{O}$ cancel, leaving
 $\text{N}_2(\text{g}) + 2 \text{H}_2(\text{g}) \rightarrow \text{N}_2\text{H}_4(\text{l})$

$$\Delta\text{H} = 534.2 \text{ kJ} + (-483.6 \text{ kJ}) = +50.6 \text{ kJ (answer A)}$$

 **Sense check:** Unburning hydrazine costs more heat than burning the hydrogen returns, so ΔH is positive: N_2H_4 stores energy above its elements. ✓





Practice 3



given (1): $2 \text{C}_3\text{H}_6(\text{g}) + 9 \text{O}_2(\text{g}) \rightarrow 6 \text{CO}_2(\text{g}) + 6 \text{H}_2\text{O}(\text{l}) \cdot \Delta\text{H} = -4116.0 \text{ kJ}$ · given (2): $\text{C}_3\text{H}_8(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 3 \text{CO}_2(\text{g}) + 4 \text{H}_2\text{O}(\text{l}) \cdot \Delta\text{H} = -2219.9 \text{ kJ}$ · given (3): $\text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}) \cdot \Delta\text{H} = -285.8 \text{ kJ}$

A refinery adds hydrogen to propene to make propane. What is ΔH , in kJ, for the target reaction?

- A. -2181.9
- B. -4563.7
- C. -123.9
- D. -447.7
- E. $+161.9$

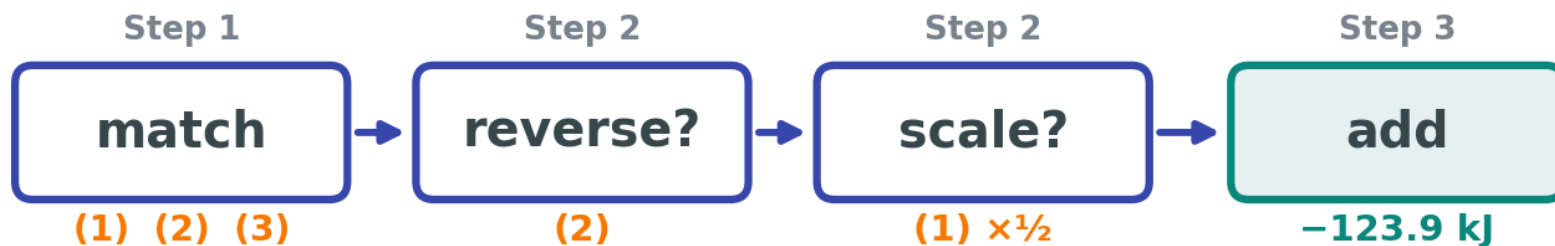
✓ Practice 3 · answer: C

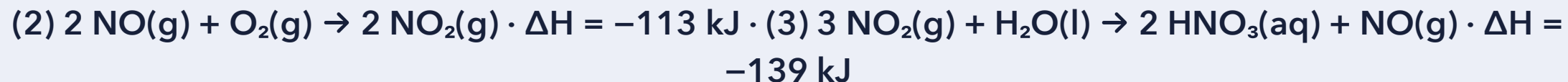
halve (1): $\text{C}_3\text{H}_6(\text{g}) + 9/2 \text{O}_2(\text{g}) \rightarrow 3 \text{CO}_2(\text{g}) + 3 \text{H}_2\text{O}(\text{l}) \cdot \Delta\text{H} = -4116.0/2 = -2058.0 \text{ kJ}$

reverse (2): $3 \text{CO}_2(\text{g}) + 4 \text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_3\text{H}_8(\text{g}) + 5 \text{O}_2(\text{g}) \cdot \Delta\text{H} = +2219.9 \text{ kJ}$ · keep (3): $-285.8 \text{ kJ} \cdot 3 \text{CO}_2, 4 \text{H}_2\text{O}, \text{ and } 5 \text{O}_2 \text{ cancel}$

$$\Delta\text{H} = -2058.0 \text{ kJ} + 2219.9 \text{ kJ} + (-285.8 \text{ kJ}) = -123.9 \text{ kJ (answer C)}$$

 **Sense check:** Propene and hydrogen burn to release 2343.8 kJ; propane burns to release 2219.9 kJ. The difference, 123.9 kJ, is released when the two combine. ✓



 Practice 4

Nitric acid is made industrially in these three steps, with all NO and NO₂ recycled. What is ΔH per mole of HNO₃ produced, in kJ/mol?

- A. -290.
- B. -381
- C. -762
- D. -1,520

✓ Practice 4 · answer: B

$$(1) + 3 \times (2) + 2 \times (3) \cdot \text{givens: } -907 \cdot -113 \cdot -139 \text{ (kJ)}$$

NO: 4 + 2 made, 6 used · NO₂: 6 made, 6 used · both cancel · 4 HNO₃ produced

$$\Delta H = -907 + 3 \times (-113) + 2 \times (-139) = -1524 \text{ kJ} \frac{-1524 \text{ kJ}}{4 \text{ mol HNO}_3}$$

$$= -381 \text{ kJ/mol (answer B)}$$

A added the givens unscaled, then divided: $-907 - 113 - 139 = -1159 \text{ kJ}$, and $-1159 \text{ kJ} \div 4 = -290. \text{ kJ/mol}$. C divided by the 2 in equation (3) instead of the 4 in the sum: -762 kJ/mol . D never divided: -1524 kJ , or $-1,520$ to three figures, is ΔH for four moles.

 **Sense check:** Tripling (2) makes the 6 NO₂ that (3), doubled, consumes. Every step is exothermic, so the per-mole answer must come out negative. ✓



Check yourself

1. A given equation must be reversed and tripled before adding. State the two changes to its ΔH .
2. Two givens sum to the target only if a species cancels. On which sides of the two equations must that species appear?

Methane, acetylene, and hydrazine were each built from their elements. Tabulating one ΔH value per compound turns Hess's law into a single formula: standard enthalpies of formation.

11 · Standard Enthalpies of Formation

Write the thermochemical equation behind a standard enthalpy of formation, assign zero to elements in their standard states, and calculate $\Delta H^{\circ}_{\text{rxn}}$ as $\sum n \cdot \Delta H^{\circ}_{\text{f}}(\text{products}) - \sum n \cdot \Delta H^{\circ}_{\text{f}}(\text{reactants})$, in either direction.

One short table, thousands of reactions

ΔH°_f (kJ/mol, 25 °C)	
H ₂ O(l)	-285.8
CO ₂ (g)	-393.5
CH ₄ (g)	-74.8
NH ₃ (g)	-46.1
O ₂ (g)	0



burn methane
-890.3 kJ
make ammonia
-92.2 kJ
any reaction from
these rows

predicted, not measured

one short table predicts the heat of thousands of reactions

Nobody measures the heat of every reaction. Each compound's formation enthalpy is measured once and tabulated, and the table then predicts ΔH for any reaction built from its rows.

Hess's law, run through the elements



target: $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$ at 25 °C · reverse (1): +285.8 kJ · add (2): +285.8 + (−241.8) = +44.0 kJ

Both equations build water from the same elements. Reverse one, add the other: the elements cancel. Evaporation absorbs heat, +44.0 kJ. Tabulate each compound's equation from its elements once, and every reaction's ΔH follows.

Standard enthalpy of formation, ΔH°_f



exactly 1 mol of product · from elements in their most stable forms · at standard conditions: 25 °C, 1 atm

ΔH°_f is the enthalpy change when one mole of a compound forms from its elements in their most stable forms at standard conditions. The product's coefficient stays 1; fractions on elements are allowed.

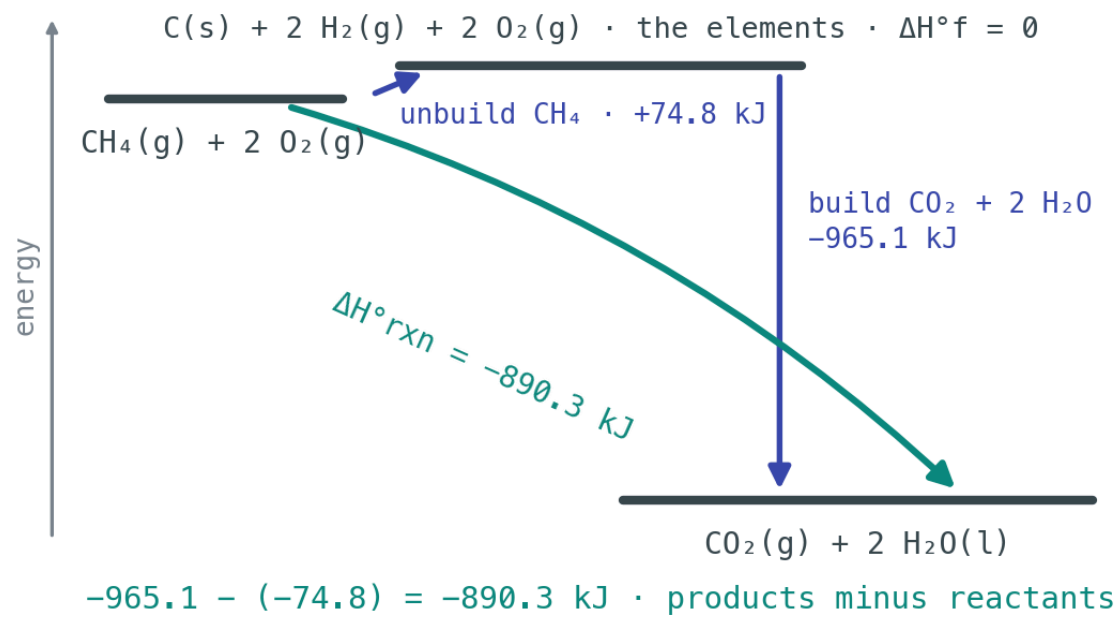
Elements in their standard states: $\Delta H^\circ_f = 0$

zero: $\text{O}_2(\text{g}) \cdot \text{N}_2(\text{g}) \cdot \text{H}_2(\text{g}) \cdot \text{Cl}_2(\text{g}) \cdot \text{Br}_2(\text{l}) \cdot \text{Hg}(\text{l}) \cdot \text{C}(\text{s, graphite}) \cdot \text{Fe}(\text{s})$

not zero: $\text{O}(\text{g}) \cdot \text{O}_3(\text{g}) \cdot \text{Br}_2(\text{g}) \cdot \text{Fe}(\text{l}) \cdot \text{C}(\text{s, diamond}) \cdot \text{every compound}$

An element already in its most stable form has nothing to form, so its ΔH°_f is zero. The stable form matters: liquid bromine, not gaseous; graphite, not diamond.

One formula replaces the cycle



$$\Delta H^\circ_{\text{rxn}} = \sum n \cdot \Delta H^\circ_f(\text{products}) - \sum n \cdot \Delta H^\circ_f(\text{reactants})$$

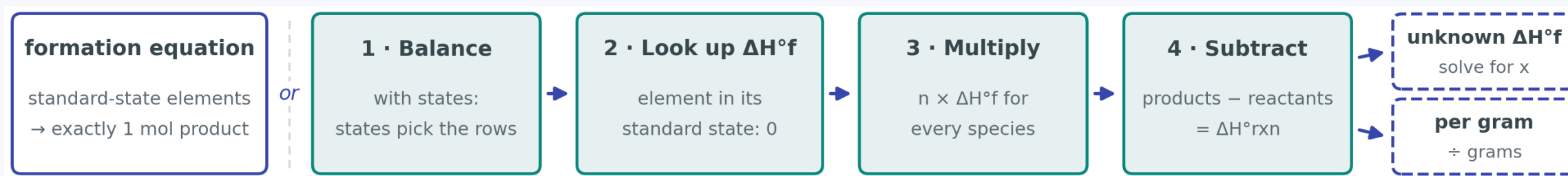
n : the balanced-equation coefficients $\cdot \Delta = \text{final} - \text{initial}$, so products come first

Every reaction can run through the elements: unbuild the reactants, rebuild the products. Adding the two legs gives products minus reactants.



The method

1. **Balance the equation with states.** The state picks the table row.
2. **Look up every ΔH°_f .** Elements in their standard states count as zero.
3. **Multiply each value by its coefficient.**
4. **Subtract: products minus reactants.**





Worked example 1: writing a formation equation

wanted: the thermochemical equation for ΔH°_f of ammonia · $\Delta H^\circ_f \text{NH}_3(\text{g}) = -46.1$
kJ/mol

reactants: elements in their most stable forms · product coefficient: 1

Write the thermochemical equation whose ΔH is the standard enthalpy of formation of $\text{NH}_3(\text{g})$.

✓ Worked example 1: solution

wanted: the formation equation of $\text{NH}_3(\text{g}) \cdot \Delta\text{H}^\circ\text{f} = -46.1 \text{ kJ/mol}$
 nitrogen's stable form: $\text{N}_2(\text{g})$ · hydrogen's stable form: $\text{H}_2(\text{g})$

Elements in their most stable forms

Exactly one mole of product



N: 1 = 1 · H: 3 = 3 · NH_3 keeps coefficient 1, so fractions land on the elements

🧠 **Sense check:** Clearing the fractions gives $\text{N}_2 + 3 \text{H}_2 \rightarrow 2 \text{NH}_3$ with $\Delta\text{H} = -92.2 \text{ kJ}$: a valid equation, but for 2 mol, so its ΔH is no longer the per-mole $\Delta\text{H}^\circ\text{f}$. ✓

formation equation

or

1 · Balance

→

2 · Look up $\Delta\text{H}^\circ\text{f}$

→

3 · Multiply

→

4 · Subtract

→

solve for x
÷ grams

Guided example: aluminum burns



table: $\text{Al(s)} \ 0 \cdot \text{O}_2\text{(g)} \ 0 \cdot \text{Al}_2\text{O}_3\text{(s)} \ -1669.8 \text{ (kJ/mol)}$

Sparklers burn aluminum powder in air. Calculate $\Delta H^\circ_{\text{rxn}}$ from the table. Run all four steps, even where a value is zero.

✓ Guided example: solution



Al: 4 = 4 · O: 6 = 6 · table: Al(s) 0 · O₂(g) 0 · Al₂O₃(s) –1669.8 (kJ/mol)

Step 1 · Balance the equation with states

Step 2 · Look up every ΔH°_f

Step 3 · Multiply each value by its coefficient

products: $2 \times (-1669.8 \text{ kJ}) = -3339.6 \text{ kJ}$ · reactants: $4 \times 0 + 3 \times 0 = 0 \text{ kJ}$

Step 4 · Subtract: products minus reactants

$\Delta H^\circ_{\text{rxn}} = -3339.6 \text{ kJ} - 0 \text{ kJ} = -3339.6 \text{ kJ}$

 **Sense check:** Only the product carries a value, so $\Delta H^\circ_{\text{rxn}}$ is $2 \times \Delta H^\circ_f$ of Al₂O₃: its formation equation, doubled. ✓



 Worked example 2: $\Delta H^\circ_{\text{rxn}}$ from the table

table: $\text{CH}_4(\text{g}) -74.8$ · $\text{O}_2(\text{g}) 0$ · $\text{CO}_2(\text{g}) -393.5$ · $\text{H}_2\text{O}(\text{l}) -285.8$ (kJ/mol)

Calculate $\Delta H^\circ_{\text{rxn}}$ for burning one mole of methane, using only the table values.

✓ Worked example 2: solution



table: $\text{CH}_4(\text{g}) -74.8$ · $\text{O}_2(\text{g}) 0$ · $\text{CO}_2(\text{g}) -393.5$ · $\text{H}_2\text{O}(\text{l}) -285.8$ (kJ/mol)

Step 1 · Balance the equation with states

Step 2 · Look up every ΔH°_f

Step 3 · Multiply each value by its coefficient

products: $-393.5 + 2 \times (-285.8) = -965.1$ kJ

reactants: $-74.8 + 2 \times 0 = -74.8$ kJ

Step 4 · Subtract: products minus reactants

$$\Delta H^\circ_{\text{rxn}} = -965.1 \text{ kJ} - (-74.8 \text{ kJ}) = -890.3 \text{ kJ}$$

 **Sense check:** Combustion is strongly exothermic. The small negative reactant sum only trims the total. ✓



Your turn: carbon monoxide burns



table: $\text{CO(g)} -110.5$ · $\text{O}_2\text{(g)} 0$ · $\text{CO}_2\text{(g)} -393.5$ (kJ/mol)

Multiply each table value by its coefficient, then subtract.

$$\Delta H^\circ_{\text{rxn}} = 2 \times (\text{ } \text{kJ}) - [2 \times (\text{ } \text{kJ}) + 0] = \text{ } \text{kJ}$$

$$\Delta H^\circ_{\text{rxn}} = 2 \times (-393.5 \text{ kJ}) - [2 \times (-110.5 \text{ kJ}) + 0] = -787.0 \text{ kJ} + 221.0 \text{ kJ} = \mathbf{-566.0 \text{ kJ}}$$

 **Sense check:** Adding carbon's two combustion steps by Hess's law gives the same -566.0 kJ : the table method is Hess's law, pre-packaged. ✓

formation equation

or

1 · Balance



2 · Look up ΔH°_f



3 · Multiply

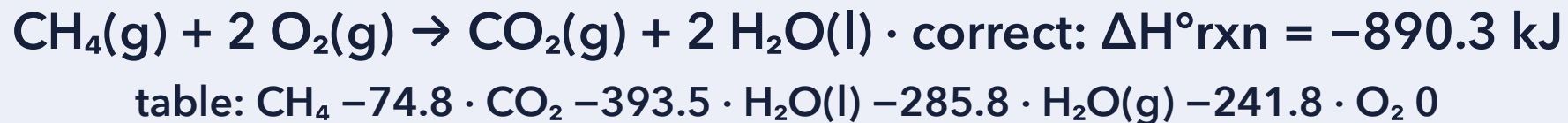


4 · Subtract



solve for x
÷ grams

⚠ Where this goes wrong



Reactants minus products. The reversed subtraction gives $+890.3 \text{ kJ}$, a combustion that absorbs heat. Products minus reactants: -890.3 kJ .

Skipping a coefficient. Counting H_2O once gives $-393.5 - 285.8 + 74.8 = -604.5 \text{ kJ}$. The 2 multiplies its value: $2 \times (-285.8)$.

Using the wrong state. With the $\text{H}_2\text{O}(\text{g})$ value the sum is $-877.1 + 74.8 = -802.3 \text{ kJ}$, the answer to a different reaction. The equation says liquid: -285.8 .

Giving an element a value. $\text{O}_2(\text{g})$ is an element in its standard state: it contributes 0 to either sum.



Practice 1

$\text{CH}_3\text{OH(l)}$: methanol $\cdot \Delta H^\circ_f = -238.7 \text{ kJ/mol}$
elements: C $\cdot$ H $\cdot$ O

Methanol fuels drag racers. Which thermochemical equation has ΔH equal to ΔH°_f of $\text{CH}_3\text{OH(l)}$?

- A. $\text{C(g)} + 4 \text{ H(g)} + \text{O(g)} \rightarrow \text{CH}_3\text{OH(l)}$
- B. $\text{CO(g)} + 2 \text{ H}_2\text{(g)} \rightarrow \text{CH}_3\text{OH(l)}$
- C. $2 \text{ C(s, graphite)} + 4 \text{ H}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2 \text{ CH}_3\text{OH(l)}$
- D. $\text{C(s, graphite)} + 2 \text{ H}_2\text{(g)} + \frac{1}{2} \text{ O}_2\text{(g)} \rightarrow \text{CH}_3\text{OH(l)}$
- E. $\text{C(s, diamond)} + 2 \text{ H}_2\text{(g)} + \frac{1}{2} \text{ O}_2\text{(g)} \rightarrow \text{CH}_3\text{OH(l)}$

✓ Practice 1 · answer: D



C: 1 = 1 · H: 4 = 4 · O: 1 = 1 · reactants: elements in their most stable forms · product: exactly 1 mol

A starts from free gaseous atoms, not C(s, graphite), H₂(g) and O₂(g). B starts from CO, a compound: industry makes methanol this way, but its ΔH is not ΔH°_f . C makes 2 mol, so its ΔH is $2 \times (-238.7) = -477.4 \text{ kJ}$. E uses diamond, not carbon's most stable form.

 **Sense check:** One mole of product from elements in their standard states. The $\frac{1}{2}$ on O₂ is allowed. ✓



 Practice 2

table: $\text{NaHCO}_3(\text{s}) -950.8 \cdot \text{Na}_2\text{CO}_3(\text{s}) -1130.7 \cdot \text{H}_2\text{O}(\text{g}) -241.8 \cdot \text{CO}_2(\text{g}) -393.5 \text{ (kJ/mol)}$

Using the table values, what is $\Delta H^\circ_{\text{rxn}}$, in kJ, for the decomposition of baking soda?

- A. -815.2
- B. -1766.0
- C. $+135.6$
- D. -135.6

✓ Practice 2 · answer: C



products: $-1130.7 + (-241.8) + (-393.5) = -1766.0 \text{ kJ}$ · reactants: $2 \times (-950.8) = -1901.6 \text{ kJ}$

$$\Delta H^\circ_{\text{rxn}} = -1766.0 \text{ kJ} - (-1901.6 \text{ kJ}) = +135.6 \text{ kJ (answer C)}$$

A dropped the coefficient 2 on NaHCO_3 : $-1766.0 - (-950.8) = -815.2$. B stopped at the products' sum: -1766.0 , before subtracting the reactants. D subtracted reactants minus products: -135.6 .

 **Sense check:** The oven must keep supplying heat for the powder to break down, so $\Delta H^\circ_{\text{rxn}}$ is positive: a formation table predicts endothermic reactions too. ✓



Worked example 3: solving for an unknown ΔH°_f



table: $\text{SO}_2(\text{g}) -296.8$ · $\text{O}_2(\text{g}) 0$ · $\text{SO}_3(\text{g})$: not listed

The reaction's $\Delta H^\circ_{\text{rxn}}$ was measured, but the table has no row for $\text{SO}_3(\text{g})$. Find ΔH°_f of $\text{SO}_3(\text{g})$.

✓ Worked example 3: solution



table: $\text{SO}_2(\text{g}) -296.8 \cdot \text{O}_2(\text{g}) 0 \cdot x = \Delta H^\circ_{\text{f}} \text{ of } \text{SO}_3(\text{g})$

Write the formula with x

$$-197.8 = 2x - [2 \times (-296.8) + 0]$$

products: $2x$ · reactants: $2 \times (-296.8) = -593.6 \text{ kJ}$

Solve for x

$$2x = -197.8 \text{ kJ} + (-593.6 \text{ kJ}) = -791.4 \text{ kJ} \quad x = -395.7 \text{ kJ/mol}$$

🧠 **Sense check:** Run it forward as a check: $2 \times (-395.7) - 2 \times (-296.8) = -197.8 \text{ kJ}$. ✓





Worked example 4: the thermite reaction



table: $\text{Al(s)} \ 0 \cdot \text{Fe}_2\text{O}_3\text{(s)} \ -822.2 \cdot \text{Al}_2\text{O}_3\text{(s)} \ -1669.8 \cdot \text{Fe(l)} \ +12.40 \text{ (kJ/mol)} \cdot \text{Al} = 26.98 \text{ g/mol}$

Aluminum powder strips the oxygen from iron oxide, leaving molten iron. Calculate the heat released in **kJ per gram of aluminum** reacted.

Two species here are elements. Tempting: write two zeros. Check each state first.

✓ Worked example 4: solution



$$\text{Al(s)} 0 \cdot \text{Fe}_2\text{O}_3\text{(s)} -822.2 \cdot \text{Al}_2\text{O}_3\text{(s)} -1669.8 \cdot \text{Fe(l)} +12.40 \text{ (kJ/mol)} \cdot \text{Al} = 26.98 \text{ g/mol}$$

Step 1 · Balance the equation with states

Step 2 · Look up every ΔH°_f

Step 3 · Multiply each value by its coefficient

$$\text{products: } -1669.8 + 2 \times (+12.40) = -1645.0 \text{ kJ}$$

$$\text{reactants: } 2 \times 0 + (-822.2) = -822.2 \text{ kJ}$$

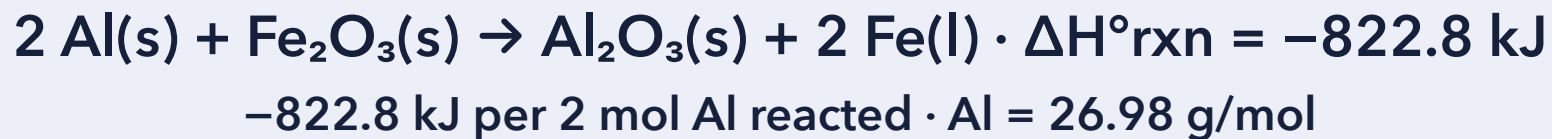
Step 4 · Subtract: products minus reactants

$$\Delta H^\circ_{\text{rxn}} = -1645.0 \text{ kJ} - (-822.2 \text{ kJ}) = -822.8 \text{ kJ}$$

 **Sense check:** Strongly exothermic, as anything that leaves its iron molten must be. ✓



✓ Worked example 4: per gram of aluminum



Convert the reaction total to a per-gram figure

$$\frac{-822.8 \text{ kJ}}{2 \text{ mol Al}} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = \frac{-822.8 \text{ kJ}}{53.96 \text{ g Al}} = -15.25 \text{ kJ/g Al}$$

 **Sense check:** Hot enough that the iron comes out molten. Run that liquid iron into a mold between two rail ends and it cools into a weld: rails are joined this way. ✓



 Practice 3

table: $\text{CO}_2(\text{g}) -393.5$ · $\text{H}_2\text{O}(\text{l}) -285.8$ (kJ/mol) · $\text{C}_2\text{H}_6(\text{g})$: not listed

A burner test measures ethane's combustion enthalpy above. What is ΔH°_f of $\text{C}_2\text{H}_6(\text{g})$, in kJ/mol?

- A. -3204.1
- B. -84.7
- C. -169.4
- D. -338.8

✓ Practice 3 · answer: B

$$-3119.4 = [4 \times (-393.5) + 6 \times (-285.8)] - [2x + 7 \times 0]$$

O₂(g) is an element in its standard state: 0 · products: $-1574.0 + (-1714.8) = -3288.8 \text{ kJ} \cdot x = \Delta H^\circ_f$ of C₂H₆(g)

$$2x = -3288.8 \text{ kJ} - (-3119.4 \text{ kJ}) = -169.4 \text{ kJ} \quad x = \mathbf{-84.7 \text{ kJ/mol (answer B)}}$$

A wrote reactants minus products: $-3119.4 = 2x - (-3288.8)$ gives $x = -3204.1$. C stopped at $2x$: -169.4 kJ belongs to 2 mol of ethane. D multiplied by the 2 instead of dividing: $2 \times (-169.4) = -338.8$.

🧠 **Sense check:** Run it forward: $4 \times (-393.5) + 6 \times (-285.8) - 2 \times (-84.7) = -3119.4 \text{ kJ}$. ✓

formation equation

or

1 · Balance



2 · Look up ΔH°_f



3 · Multiply



4 · Subtract



solve for x

÷ grams

 Practice 4

table: $\text{C}_5\text{H}_{12}(\text{l}) -146.8 \cdot \text{CO}_2(\text{g}) -393.5 \cdot \text{H}_2\text{O}(\text{g}) -241.8 \cdot \text{H}_2\text{O}(\text{l}) -285.8 \cdot \text{O}_2(\text{g}) 0$ (kJ/mol)

Pentane burns in a camp stove. What is $\Delta H^\circ_{\text{rxn}}$, in kJ?

- A. -3535.5
- B. -3418.3
- C. -3271.5
- D. $+3271.5$

✓ Practice 4 · answer: C



products: $5 \times (-393.5) + 6 \times (-241.8) = -3418.3 \text{ kJ}$ · reactants: $-146.8 + 8 \times 0 = -146.8 \text{ kJ}$

$$\Delta H^\circ_{\text{rxn}} = -3418.3 \text{ kJ} - (-146.8 \text{ kJ}) = \mathbf{-3271.5 \text{ kJ (answer C)}}$$

A took the liquid-water row against the equation's (g): -3535.5 . B stopped at the products' sum: $5 \times (-393.5) + 6 \times (-241.8) = -3418.3$, before subtracting the reactant. D subtracted reactants minus products: $+3271.5$.

 **Sense check:** Five carbons of fuel against methane's one, so several times methane's 890 kJ , and negative as every combustion must be. ✓





Check yourself

1. State the two conditions a species must meet for its ΔH°_f to equal zero.
2. A reaction's $\Delta H^\circ_{\text{rxn}}$ and every ΔH°_f except one are known. Outline how to find the missing value.

The table predicts ΔH from tabulated wholes; bond energies tell the same story from inside the molecule, where breaking and making individual bonds sets the size and sign of ΔH .

Can you...?

- ☐ distinguish heat from work and convert among joules, calories, and Calories?
- ☐ identify exothermic and endothermic processes from equations, energy diagrams, and the sign of ΔH ?
- ☐ apply $q = m \cdot c \cdot \Delta T$ to find heat, mass, specific heat, or temperature change?
- ☐ compare how specific heats set different substances' temperature response to the same heat?
- ☐ use calorimetry data to determine the heat of a process or reaction?
- ☐ reverse and scale thermochemical equations and use ΔH as a conversion factor?
- ☐ combine reversed and scaled equations with Hess's law to find an overall reaction's enthalpy?
- ☐ write formation equations and use standard formation enthalpies to find any reaction's enthalpy?
- ☐ compute heat for melting, vaporizing, or a full heating curve?

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