

Enthalpy and Calorimetry

What is enthalpy and how do you measure the heat of a reaction?

●●○ Intermediate

Thermochemistry & Thermodynamics

Last reviewed 4 October 2026

Before you start: it helps to understand [The Mole Concept](#) and [Temperature Scales: Celsius, Fahrenheit and Kelvin](#).

What is it?

Chemical reactions almost always **release or absorb heat**. The heat change of a reaction at constant pressure is its **enthalpy change**, ΔH , usually given in **kJ/mol**.

- **Exothermic** reaction: heat is **released** to the surroundings, which warm up. ΔH is **negative**. Examples: combustion, neutralization, respiration.
- **Endothermic** reaction: heat is **absorbed** from the surroundings, which cool down. ΔH is **positive**. Examples: dissolving ammonium nitrate (instant cold packs), photosynthesis, thermal decomposition.

💡 KEY IDEA

$\Delta H = H_{\text{products}} - H_{\text{reactants}}$. Negative: exothermic (heat out). Positive: endothermic (heat in). The sign is always from the point of view of the **reaction** (the system).

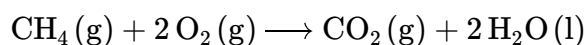
Why does it matter?

- **Energy.** Fuels are compared by the heat their combustion releases per gram or per mole.
- **Food.** The energy content of food (in kJ or kcal) is measured by burning it in a calorimeter.
- **Safety and design.** Chemists must know how much heat a reaction gives out to cool it safely, or how much to supply to keep it going.

How does it work?

1. Thermochemical equations

A thermochemical equation gives ΔH for the amounts in the **balanced equation**:



$$\Delta H = -890.5 \text{ kJ}$$

- Doubling the equation **doubles** ΔH .

- Reversing the equation **changes the sign** of ΔH .
- **States matter:** forming $\text{H}_2\text{O}(\text{g})$ instead of $\text{H}_2\text{O}(\text{l})$ releases less heat.

2. Specific heat capacity and $q = mc\Delta T$

The **specific heat capacity**, c , is the heat needed to raise the temperature of **1 g** of a substance by **1 °C** (or 1 K). For water, $c = 4.184 \text{ J}/(\text{g}\cdot^\circ\text{C})$. The heat q gained or lost is:

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

where m is the mass in g and $\Delta T = T_{\text{final}} - T_{\text{initial}}$. A change of 1 °C equals a change of 1 K, so ΔT is the same on both scales.

3. Calorimetry

A **calorimeter** measures heat changes. In a simple **coffee-cup calorimeter**, a reaction happens in water inside insulated cups:

1. Measure the temperature change of the solution.
2. Calculate $q_{\text{solution}} = mc\Delta T$ (assume the solution behaves like water: density 1.00 g/mL, $c = 4.184 \text{ J}/(\text{g}\cdot^\circ\text{C})$).
3. The reaction's heat is the opposite: $q_{\text{reaction}} = -q_{\text{solution}}$ (heat lost by one is gained by the other).
4. Divide by the moles that reacted: $\Delta H = \frac{q_{\text{reaction}}}{n}$.

THINK OF IT LIKE THIS

Think of the reaction as a heater placed in a bath. You can't see the heat directly, but the bath's thermometer tells you: if the water warms, the "heater" gave heat out (exothermic); if the water cools, the reaction took heat in (endothermic).

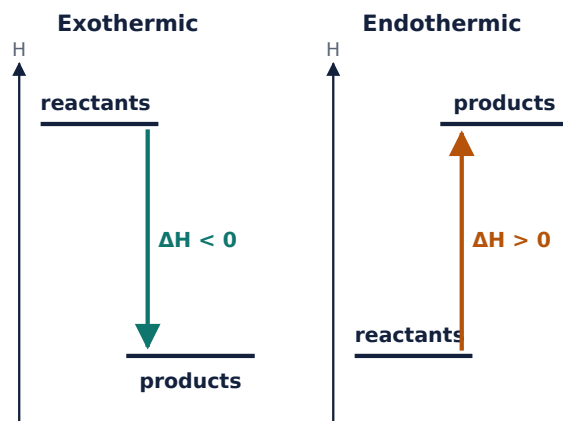
MORE PRECISELY

Enthalpy, H , is defined as $H = U + PV$; at constant pressure, the heat transferred equals ΔH . Coffee-cup calorimeters work at constant (atmospheric) pressure, so they measure ΔH directly. A **bomb calorimeter** works at constant volume, measures ΔU , and is used for combustion. Simple calorimeters lose some heat to the surroundings and the cups, so measured values are usually slightly smaller in size than accepted ones.

Visualise it

Exothermic and endothermic reactions

$$\Delta H = H(\text{products}) - H(\text{reactants})$$



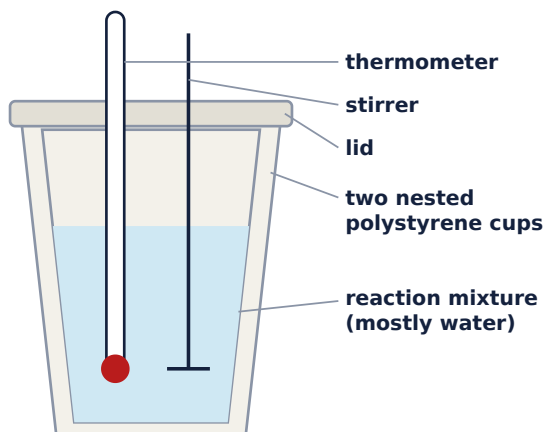
Exothermic: heat released to the surroundings (they warm up)

Endothermic: heat absorbed from the surroundings (they cool)

ΔH is negative when the products are lower in enthalpy than the reactants.

A coffee-cup calorimeter

The heat of reaction changes the temperature of the solution.



$$q(\text{solution}) = m \times c \times \Delta T; q(\text{reaction}) = -q(\text{solution})$$

The insulated cups keep heat from escaping (constant pressure)

The solution's temperature change measures the heat of the reaction.

Worked example

WORKED EXAMPLE: Heating water

Question: How much heat is needed to warm 250. g of water from 20.0 °C to 80.0 °C?

Step 1 $\Delta T = 80.0\text{ }^{\circ}\text{C} - 20.0\text{ }^{\circ}\text{C} = 60.0\text{ }^{\circ}\text{C}$

Step 2 Substitute:

$$\begin{aligned} q &= mc\Delta T \\ &= (250.\text{ g}) \\ &\quad \times \left(4.184\text{ }\frac{\text{J}}{\text{g}\cdot^{\circ}\text{C}}\right) \\ &\quad \times (60.0\text{ }^{\circ}\text{C}) \\ &= 6.28 \times 10^4\text{ J} = 62.8\text{ kJ} \end{aligned}$$

Step 3 The units of g and °C cancel, leaving J.

 WORKED EXAMPLE: Enthalpy of neutralization

Question: 50.0 mL of 1.00 M HCl and 50.0 mL of 1.00 M NaOH, both at 21.0 °C, are mixed in a coffee-cup calorimeter. The temperature rises to 27.8 °C. Calculate ΔH per mole of water formed.

Step 1 Mass of solution: $100.0 \text{ mL} \times 1.00 \text{ g/mL} = 100. \text{ g}$; $\Delta T = 27.8 \text{ }^{\circ}\text{C} - 21.0 \text{ }^{\circ}\text{C} = 6.8 \text{ }^{\circ}\text{C}$

Step 2 Heat gained by the solution:

$$\begin{aligned} q_{\text{soln}} &= (100. \text{ g}) \\ &\times \left(4.184 \frac{\text{J}}{\text{g}\cdot^{\circ}\text{C}} \right) \\ &\times (6.8 \text{ }^{\circ}\text{C}) \\ &= 2845 \text{ J} \end{aligned}$$

Step 3 $q_{\text{reaction}} = -2845 \text{ J}$ (the reaction released this heat).

Step 4 Moles of water formed: $n = 1.00 \text{ mol/L} \times 0.0500 \text{ L} = 0.0500 \text{ mol}$

Step 5 Divide:

$$\begin{aligned} \Delta H &= \frac{-2845 \text{ J}}{0.0500 \text{ mol}} \\ &= -5.7 \times 10^4 \text{ J/mol} \\ &= -57 \text{ kJ/mol} \end{aligned}$$

Step 6 ΔT (6.8 °C) has 2 significant figures, so the answer has 2. It is negative: the reaction is **exothermic**.

Common mistake

 COMMON MISTAKE: Getting the sign wrong

If the solution warms up, the reaction **gave out** heat: q_{reaction} and ΔH are **negative**, even though q_{solution} is positive.

 COMMON MISTAKE: Using only the mass of the solid

In $q = mc\Delta T$, m is the mass of the **solution that changes temperature** (mostly water), not the mass of the substance that reacted.

 COMMON MISTAKE: Forgetting to divide by moles

q is the heat for **this** experiment (in J). ΔH is per **mole** (kJ/mol), so divide by the moles that reacted and convert J to kJ.

 NOTATION NOTE

- ΔH° (with a degree sign) is a **standard** enthalpy change: 1 bar and a stated temperature, usually 25 °C.
- Units: c in $\text{J}/(\text{g}\cdot^\circ\text{C}) = \text{J}/(\text{g}\cdot\text{K}) = \text{J g}^{-1} \text{K}^{-1}$; ΔH in $\text{kJ/mol} = \text{kJ mol}^{-1}$.
- Some textbooks write $q = mc\Delta T$ as $q = mc_s\Delta T$, with c_s for specific heat.

Remember this

 REMEMBER THIS

- **Exothermic:** heat released, ΔH negative. **Endothermic:** heat absorbed, ΔH positive.
- $q = mc\Delta T$; for water $c = 4.184 \text{ J}/(\text{g}\cdot^\circ\text{C})$; ΔT in °C equals ΔT in K.
- **Calorimetry:** $q_{\text{reaction}} = -q_{\text{solution}}$; $\Delta H = q_{\text{reaction}}/n$, in kJ/mol.
- Doubling an equation doubles ΔH ; reversing it changes the sign.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

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