

Phase Changes and Heating Curves

What happens to energy and temperature when a substance melts or boils?

●●○ Intermediate

States of Matter & Gases

Last reviewed 5 October 2026

Before you start: it helps to understand [Intermolecular Forces](#) and [Enthalpy and Calorimetry](#).

What is it?

A **phase change** (change of state) turns a substance from one state into another without changing what it is: ice, liquid water and steam are all H_2O .

Change	From → to	Energy
Melting (fusion)	solid → liquid	absorbed (endothermic)
Freezing	liquid → solid	released (exothermic)
Vaporization (boiling, evaporation)	liquid → gas	absorbed
Condensation	gas → liquid	released
Sublimation	solid → gas	absorbed
Deposition	gas → solid	released

Changes that pull particles **apart** (overcoming intermolecular forces) absorb energy; changes that bring them **together** release it.

KEY IDEA

While a pure substance is melting or boiling, its **temperature stays constant**, even though heat is still being added. All the energy goes into overcoming the forces between particles, not into making them move faster. That's why a heating curve has **flat sections**.

Why does it matter?

- **Cooling and heating.** Sweat cools you because evaporation absorbs heat; ice keeps drinks at $0\text{ }^{\circ}\text{C}$ while it melts; refrigerators work by evaporating and condensing a fluid.
- **Steam burns.** Steam at $100\text{ }^{\circ}\text{C}$ burns far worse than water at $100\text{ }^{\circ}\text{C}$, because it releases its large enthalpy of vaporization as it condenses on the skin.

- **Cooking and climate.** Water boils at a lower temperature at altitude; huge amounts of energy are moved around the planet as water evaporates and condenses.

How does it work?

1. The heating curve

Heating ice at $-15\text{ }^{\circ}\text{C}$ steadily until it becomes steam at $125\text{ }^{\circ}\text{C}$ gives five stages:

1. ice warms (sloped): $q = mc_{\text{ice}}\Delta T$
2. ice melts at $0\text{ }^{\circ}\text{C}$ (flat): $q = n\Delta H_{\text{fus}}$
3. water warms (sloped): $q = mc_{\text{water}}\Delta T$
4. water boils at $100\text{ }^{\circ}\text{C}$ (flat): $q = n\Delta H_{\text{vap}}$
5. steam warms (sloped): $q = mc_{\text{steam}}\Delta T$

The total heat is the sum of the five steps.

2. Enthalpies of fusion and vaporization

The **enthalpy of fusion**, ΔH_{fus} , is the energy to melt one mole of solid at its melting point; the **enthalpy of vaporization**, ΔH_{vap} , is the energy to vaporize one mole of liquid. For water, $\Delta H_{\text{fus}} = 6.01\text{ kJ/mol}$ and $\Delta H_{\text{vap}} = 40.7\text{ kJ/mol}$. Vaporization needs far more energy because the molecules must be separated **completely**, while melting only loosens them.

3. Vapour pressure and boiling

Even below its boiling point, some molecules at the surface of a liquid have enough energy to escape: **evaporation**. In a closed container, the vapour builds up until evaporation and condensation balance. The pressure of the vapour then is the **vapour pressure**, which rises steeply with temperature.

A liquid **boils** when its vapour pressure equals the **external pressure**: bubbles of vapour can then form throughout the liquid. At 1 atm (760 mmHg) this is the **normal boiling point**: $100\text{ }^{\circ}\text{C}$ for water. Where the air pressure is lower, such as on a mountain, water boils below $100\text{ }^{\circ}\text{C}$; in a pressure cooker it boils above $100\text{ }^{\circ}\text{C}$.

THINK OF IT LIKE THIS

Melting is like a crowd leaving a packed hall. While people are still squeezing out of the doors (the phase change), the crowd's speed doesn't increase: all the effort goes into getting through. Only once everyone is outside can they start to move faster (the temperature rises again).

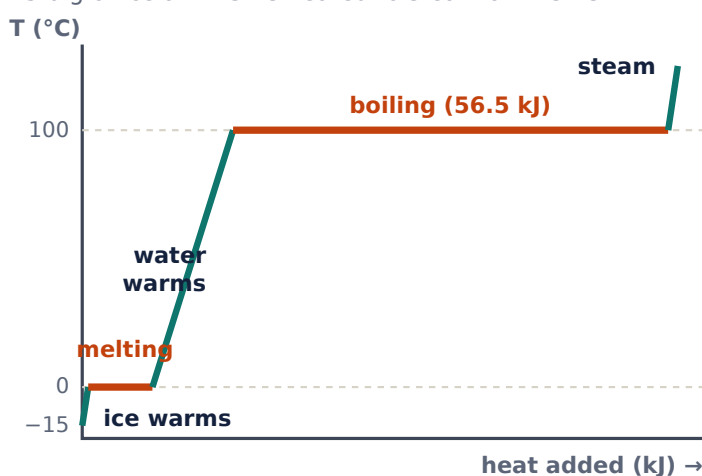
Q MORE PRECISELY

Evaporation happens at any temperature, from the surface only; boiling happens at one temperature (for a given pressure), throughout the liquid. The temperature stays constant during a phase change only while both phases are present and the pressure is constant. ΔH_{vap} itself changes slightly with temperature (water: 44.0 kJ/mol at 25 °C, 40.7 kJ/mol at 100 °C). The relationship between vapour pressure and temperature is described by the Clausius–Clapeyron equation.

Visualise it

Heating curve of water

25.0 g of ice at $-15\text{ }^{\circ}\text{C}$ heated to steam at $125\text{ }^{\circ}\text{C}$



Heating curve of water. The flat sections are phase changes; boiling takes by far the most energy.

Worked example

WORKED EXAMPLE: Melting ice

Question: How much heat is needed to melt 100. g of ice at $0\text{ }^{\circ}\text{C}$?

Step 1 Moles: $n = \frac{100.\text{ g}}{18.02\text{ g/mol}} = 5.549\text{ mol}$

Step 2 Heat:

$$\begin{aligned} q &= n\Delta H_{\text{fus}} = 5.549\text{ mol} \\ &\times 6.01\frac{\text{kJ}}{\text{mol}} \\ &= 33.4\text{ kJ} \end{aligned}$$

 WORKED EXAMPLE: A full heating curve

Question: How much heat turns 25.0 g of ice at $-15.0\text{ }^{\circ}\text{C}$ into steam at $125.0\text{ }^{\circ}\text{C}$? ($n = 25.0\text{ g} \div 18.02\text{ g/mol} = 1.387\text{ mol}$)

Step 1 Warm ice:

$$\begin{aligned}q_1 &= 25.0\text{ g} \times 2.09\frac{\text{J}}{\text{g}^{\circ}\text{C}} \\&\quad \times 15.0\text{ }^{\circ}\text{C} \\&= 784\text{ J} = 0.784\text{ kJ}\end{aligned}$$

Step 2 Melt: $q_2 = 1.387\text{ mol} \times 6.01\text{ kJ/mol} = 8.34\text{ kJ}$

Step 3 Warm water:

$$\begin{aligned}q_3 &= 25.0\text{ g} \times 4.184\frac{\text{J}}{\text{g}^{\circ}\text{C}} \\&\quad \times 100.0\text{ }^{\circ}\text{C} \\&= 10\,460\text{ J} = 10.46\text{ kJ}\end{aligned}$$

Step 4 Boil: $q_4 = 1.387\text{ mol} \times 40.7\text{ kJ/mol} = 56.47\text{ kJ}$

Step 5 Warm steam:

$$\begin{aligned}q_5 &= 25.0\text{ g} \times 2.01\frac{\text{J}}{\text{g}^{\circ}\text{C}} \\&\quad \times 25.0\text{ }^{\circ}\text{C} \\&= 1256\text{ J} = 1.26\text{ kJ}\end{aligned}$$

Step 6 Total:

$$\begin{aligned}q &= (0.784 + 8.34 + 10.46 \\&\quad + 56.47 + 1.26)\text{ kJ} \\&= 77.3\text{ kJ}\end{aligned}$$

Boiling alone is about 73 % of the total.

 WORKED EXAMPLE: Why steam burns

Question: Compare the heat released to the skin (at 37 °C) by 5.00 g of steam at 100 °C and by 5.00 g of water at 100 °C.

Step 1 Water cooling from 100 °C to 37 °C:

$$\begin{aligned} q &= 5.00 \text{ g} \times 4.184 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \\ &\quad \times 63 ^\circ\text{C} \\ &= 1318 \text{ J} = 1.32 \text{ kJ} \end{aligned}$$

Step 2 Steam first condenses: $\frac{5.00 \text{ g}}{18.02 \text{ g/mol}} \times 40.7 \text{ kJ/mol} = 11.29 \text{ kJ}$, then cools like the water (1.32 kJ): **12.6 kJ** in total.

Step 3 The steam delivers almost **ten times** as much heat.

Common mistake

 **COMMON MISTAKE:** Using $q = mc\Delta T$ during a phase change

During melting or boiling, $\Delta T = 0$, but heat is still absorbed. Use $q = n\Delta H$ for the flat parts and $q = mc\Delta T$ only for the sloped parts, with the right c for each state.

 **COMMON MISTAKE:** Mixing J and kJ

$mc\Delta T$ usually gives joules, while $n\Delta H$ gives kilojoules. Convert to the same unit before adding.

 **COMMON MISTAKE:** Thinking boiling breaks bonds within molecules

When water boils, the bubbles contain H_2O molecules. Only intermolecular forces are overcome.

 NOTATION NOTE

- “Latent heat” is an older name for the enthalpy of a phase change.
- $\Delta H(\text{fus})$ is positive for melting; for freezing the same value is negative.

Remember this

REMEMBER THIS

- Melting, vaporization, sublimation: endothermic. Freezing, condensation, deposition: exothermic.
- Temperature is constant during a phase change: all the energy overcomes intermolecular forces.
- Sloped parts: $q = mc\Delta T$. Flat parts: $q = n\Delta H_{\text{fus}}$ or $n\Delta H_{\text{vap}}$.
- Water: $\Delta H_{\text{fus}} = 6.01 \text{ kJ/mol}$; $\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$.
- A liquid boils when its vapour pressure equals the external pressure.

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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