

Temperature and Thermal Energy

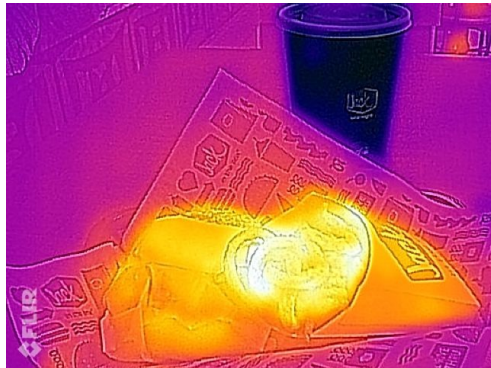
A2 Physics ■ Topic 14

A-Level Physics



What we will cover

- 1 Thermal equilibrium
- 2 Measuring temperature
- 3 The kelvin & absolute zero
- 4 Specific heat capacity
- 5 Specific latent heat
- 6 Recap

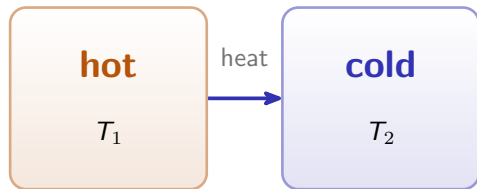




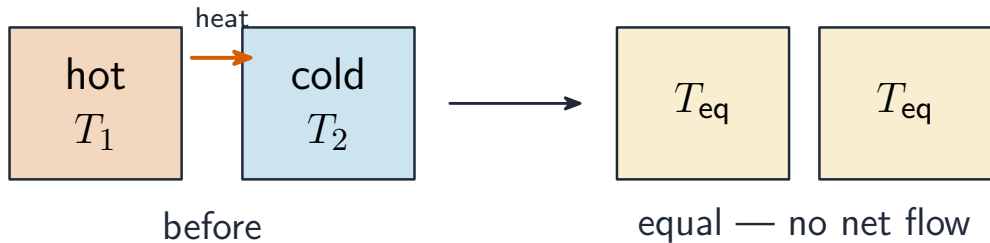
Thermal equilibrium

Two bodies in contact exchange energy until they reach the same **temperature** – **thermal equilibrium**. Heat flows from the hotter to the colder. Once they match, the net flow is zero.

A thermometer works because it comes to equilibrium with whatever it touches – so it must be small enough not to change that temperature.



What equilibrium means



Net energy flows from hot to cold until the two are at the **same temperature** – that is what a thermometer waits for before it reads.

At **thermal equilibrium** there is **no net** flow of energy between the bodies – they are at the same temperature, not necessarily holding the same energy.

Temperature is not energy

Temperature fixes the **direction** of heat flow.
It does **not** say how much **thermal energy**
a body holds.

A cup of water at 100°C holds far
less energy than a pool at 25°C – but
metal dropped in the cup **gains** energy,
and in the pool **loses** it.



hotter,
less energy



cooler,
far more energy

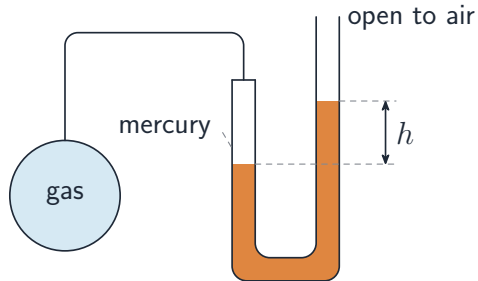


Making a thermometer

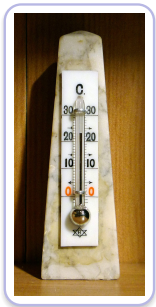
Any property that changes **repeatably** with temperature will do:

- liquid **density** – liquid-in-glass
- gas volume (or pressure) – gas thermometer
- **resistance** of a metal
- **thermocouple** e.m.f.

Two thermometers agree **only at the calibration points** – in between, each follows its own property's curve.



A thermometric property



Any property that varies reproducibly with temperature will do – here the length of a liquid column.

Choosing a thermometric property

what is needed

a property that varies **measurably and reproducibly** with **thermodynamic temperature**

liquid-in-glass

the length of a mercury or alcohol thread in a **capillary** tube

resistance

the resistance of a wire or a thermistor – non-linear, but very sensitive

a thermocouple

two different metals joined at two points; the **electromotive force** depends on the **temperature difference** between the junctions

why a thermocouple

small, fast-responding, and it works over a very wide range

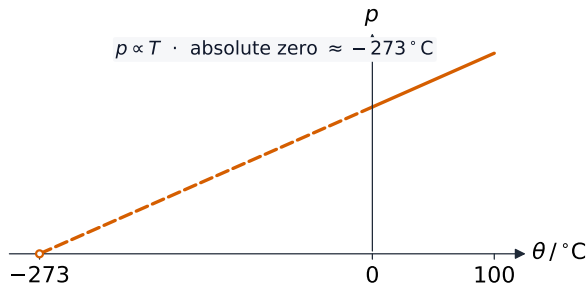
gas thermometer

the pressure of a fixed volume – the most accurate, and the standard the others are calibrated against

Sensitivity, range and response time are the three criteria. Naming which one your choice wins on is the mark.



Absolute zero



The **thermodynamic scale** leans on no single substance – only on the laws of thermodynamics. Its unit is the **kelvin**. At **absolute zero** (0 K) a system holds its least possible **internal energy**.

Cool *any* gas and extrapolate its p - θ line to zero pressure: every gas points at the **same** -273°C .

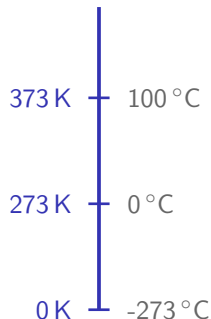
Kelvin and Celsius

$$T/\text{K} = \theta/^{\circ}\text{C} + 273.15$$

- $0^{\circ}\text{C} = 273.15\text{ K}$; $100^{\circ}\text{C} = 373.15\text{ K}$
- a kelvin and a degree Celsius are the **same size**; a kelvin reading is an **absolute temperature**
- so a **difference** of 1 K is 1°C

Gas laws and energy equations need **kelvin**. Using $^{\circ}\text{C}$ there is simply wrong – but for a ΔT either unit works.

same size steps





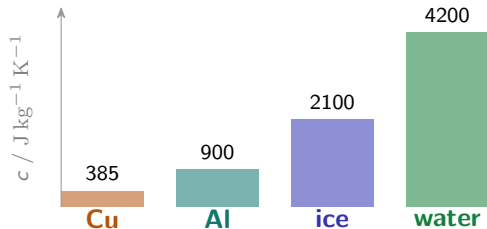
Specific heat capacity

Specific heat capacity is the energy to raise **unit mass** by **one kelvin**:

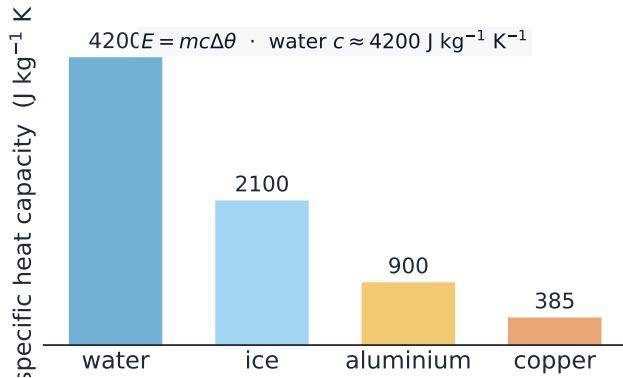
$$c = \frac{Q}{m\Delta T} \iff Q = mc\Delta T$$

$$[c] = \text{J kg}^{-1} \text{K}^{-1}.$$

Water's $c \approx 4200$ is far above any solid here – which is why it is such a good coolant, and why the oceans steady the climate.



Water's specific heat capacity is far higher



Water's specific heat capacity is far higher than common solids – why it is such a good coolant

Worked example – heating water

Heat 0.50 kg of water from 20 to 100 °C.
($c = 4200 \text{ J kg}^{-1} \text{ K}^{-1}$)

A **difference** is the same in K and °C, so
 $\Delta T = 80$:

$$Q = mc\Delta T = 0.50 \times 4200 \times 80$$
$$= 1.68 \times 10^5 \text{ J} = 168 \text{ kJ}$$

Do **not** convert 20 and 100 to kelvin first – the 273.15 cancels in the subtraction. Convert only where an **absolute** T is needed.

Worked example – mixing

Mix 0.20 kg of water at 80 °C with 0.30 kg at 20 °C, no heat lost. Find the final temperature.

heat lost by the hot = heat gained by the cold
(c cancels):

$$0.20 (80 - T_{\text{eq}}) = 0.30 (T_{\text{eq}} - 20)$$

$$\Rightarrow T_{\text{eq}} = 44^{\circ}\text{C}$$

Write both sides as a **positive** energy: (hot – T_{eq}) and (T_{eq} – cold). The answer sits nearer the **bigger** mass – here 44, not 50.



Specific latent heat

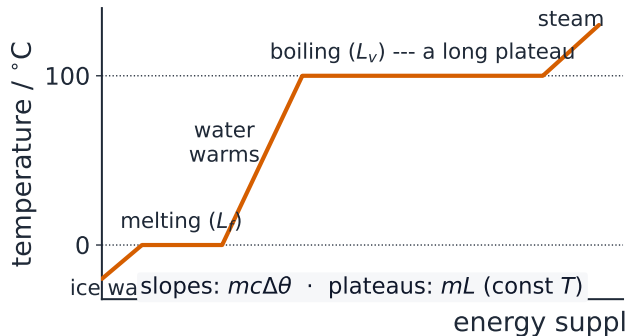
Changing state takes energy at **constant temperature** – the **latent heat** :

$$L = \frac{Q}{m} \iff Q = mL, \quad [L] = \text{J kg}^{-1}$$

- **specific latent heat of fusion** L_f : solid $\leftrightarrow$ liquid
- **specific latent heat of vapourisation** L_v : liquid $\leftrightarrow$ gas

Sloped = warming ($mc\Delta T$); **flat** = changing state (mL). Never mix the two.

Specific latent heat, in detail



Changing state takes energy at **constant temperature** – the **latent heat** :

$$L = \frac{Q}{m} \iff Q = mL, \quad [L] = \text{J kg}^{-1}$$

Why $L_v > L_f$

Bonds

melting breaks *some* bonds
– the particles stay close.
Boiling breaks *all* of them.

Work on the atmosphere

vapour has $\sim 10^3 \times$ the volume, so
it pushes the air back as it forms.

For water $L_v \approx 2.26 \times 10^6$ against $L_f \approx 3.34 \times 10^5 \text{ J kg}^{-1}$ – about **7 times** larger.

Worked example – boiling with a heater

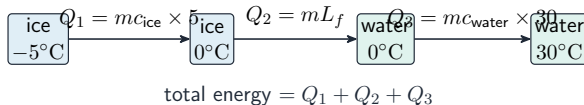
A 2.0 kW heater boils water already at 100 °C. How long to turn 0.10 kg into steam? ($L_v = 2.26 \times 10^6 \text{ J kg}^{-1}$)

$$Q = mL_v = 0.10 \times 2.26 \times 10^6 = 2.26 \times 10^5 \text{ J}$$

$$t = \frac{Q}{P} = \frac{2.26 \times 10^5}{2000} \approx 110 \text{ s}$$

Already **at** 100 °C, so there is **no** $mc\Delta T$ term – all the energy goes into the phase change.

Multi-step problems



Ice at $-5^{\circ}\text{C} \rightarrow$ water at 30°C is **three** stages:

- $Q_1 = mc_{\text{ice}} \times 5$ – warm the solid
- $Q_2 = mL_f$ – melt at 0°C
- $Q_3 = mc_{\text{water}} \times 30$ – warm the liquid

Add $Q_1 + Q_2 + Q_3$. Each phase change is a **separate** mL step at constant temperature – and the c of ice is not the c of water.

Finding c by experiment, and why boiling costs more

the method

supply known energy electrically: $Q = VIt$, from the **power** $P = VI$

then

measure the temperature rise and use $Q = mc\Delta\theta$

insulation

lagging the container reduces heat loss, which is the main source of error

if P is given

use $Q = Pt$ and assume no loss – the question usually says so

why $L_v > L_f$

melting only loosens the **intermolecular** bonds; boiling **breaks them completely**

and

the vapour must also **push back the atmosphere** as it expands, and that work is paid for too

Both reasons come from the particle picture. Giving only “the bonds break” misses the expansion work, which is the second marked point.

Exam tips

- **Thermal equilibrium** means no net heat flow (equal temperature); the thermodynamic scale is fixed by absolute zero and the triple point.
- Convert temperatures with $T/\text{K} = \theta/^{\circ}\text{C} + 273.15$ and use **kelvin** in energy and gas equations.
- Use $Q = mc\Delta T$ for heating and $Q = mL$ for a change of state (no temperature change) – never mix the two.



Topic 14 – key takeaways

1 Equilibrium

equal temperature = no *net* heat flow

2 Kelvin

$T/\text{K} = \theta/^{\circ}\text{C} + 273.15$; use K in gas laws

3 Heating

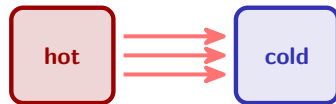
$Q = mc\Delta T$ – temperature changes

4 Changing state

$Q = mL$ – temperature does *not* change

Check yourself

- 1 Convert 27°C to kelvin.
- 2 Which equation for melting ice at 0°C ?
- 3 Why is a flat part of a heating curve flat?
- 4 Does the hotter body always hold more energy?



Answers 1. 300.15 K 2. $Q = mL_f$ 3. energy breaks bonds, not raising T 4. no