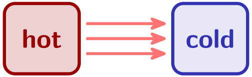
Temperature and Thermal Energy

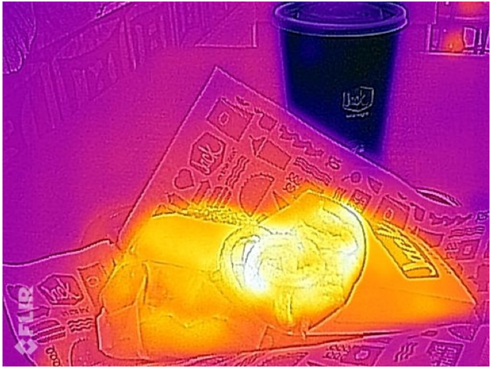
A2 Physics ▪ Topic 14

A-Level Physics

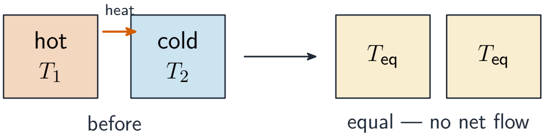


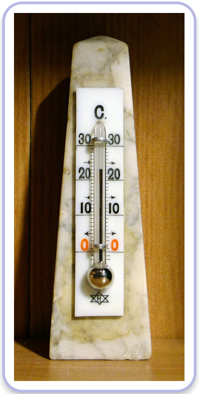
What we will cover

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



Thermal equilibrium



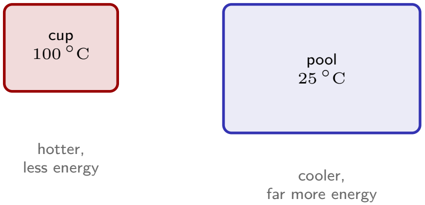


thermometer

Temperature is not energy

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

> A cup of water at $100\,^{\circ}\mathrm{C}$ holds far less energy than a pool at $25\,^{\circ}\mathrm{C}$ – but metal dropped in the cup gains energy, and in the pool loses it.

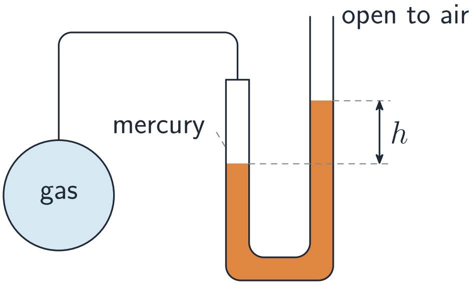


Making a thermometer

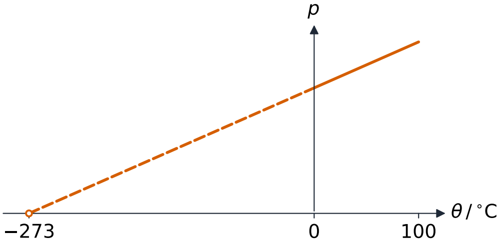
Any property that changes repeatably with temperature will do:

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

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



Absolute zero



The **thermodynamic scale** 热力学温度 leans on no single substance – only on the laws of thermodynamics. Its unit is the **kelvin** 开尔文. At **absolute zero** 绝对零度 $(0\,\mathrm{K})$ a system holds its least possible **internal energy** 内能.

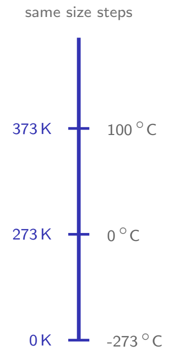
> Cool *any* gas and extrapolate its p–θ line to zero pressure: every gas points at the same $-273\,°\mathrm{C}$.

Kelvin and Celsius

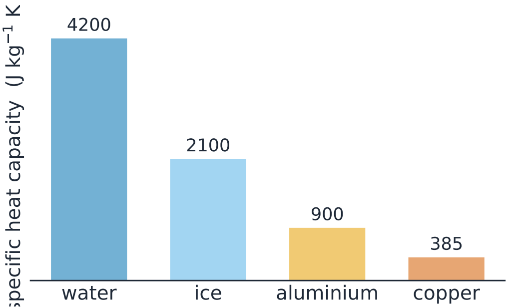
$$T/\,\mathrm{K} = \theta/\,^{\circ}\mathrm{C} + 273.15$$

- $0\,^{\circ}\mathrm{C} = 273.15\,\mathrm{K}$; $100\,^{\circ}\mathrm{C} = 373.15\,\mathrm{K}$
- a kelvin and a degree Celsius are the same size
- so a difference of $1\,\mathrm{K}$ *is* $1\,^{\circ}\mathrm{C}$

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



Specific heat capacity



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

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

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

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.

Worked example – heating water

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

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

$Q = mc\Delta T = 0.50 \times 4200 \times 80$

$= 1.68 \times 10^{5}\,\mathrm{J} = 168\,\mathrm{kJ}$

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 $0.20\,\mathrm{kg}$ of water at $80\,^{\circ}\mathrm{C}$ with $0.30\,\mathrm{kg}$ at $20\,^{\circ}\mathrm{C}$, no heat lost. Find the final temperature.

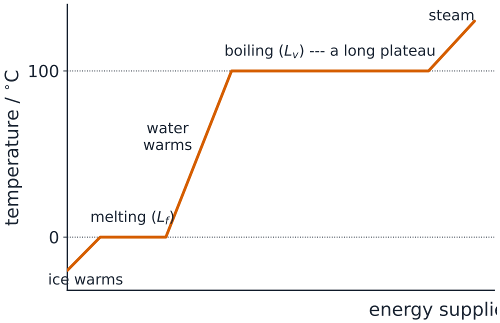
heat lost by the hot = heat gained by the cold (c cancels):

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

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

Write both sides as a positive energy: (hot $-\ T_{\mathrm{eq}}$) 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** 潜热:

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

- **fusion** 熔化 L_f: solid ↔ liquid
- **vaporisation** 汽化 L_v: liquid ↔ gas

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

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 $L_f \approx 3.34 \times 10^5\,\mathrm{J\,kg^{-1}}$ – about 7 times larger.

Worked example – boiling with a heater

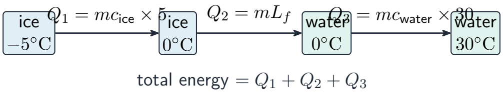
A $2.0\,\mathrm{kW}$ heater boils water already at $100\,^\circ\mathrm{C}$. How long to turn $0.10\,\mathrm{kg}$ into steam? ($L_\mathrm{v} = 2.26 \times 10^6\,\mathrm{J\,kg^{-1}}$)

$$Q = mL_\mathrm{v} = 0.10 \times 2.26 \times 10^6 = 2.26 \times 10^5\,\mathrm{J}$$

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

Already at $100\,^\circ\mathrm{C}$, so there is no $mc\Delta T$ term – all the energy goes into the phase change.

Multi-step problems



Ice at $-5\,°C$ → water at $30\,°C$ is three stages:

- $Q_1 = mc_{\text{ice}} \times 5$ – warm the solid
- $Q_2 = mL_{\text{f}}$ – 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.

Topic 14 – key takeaways

1 Equilibrium

equal temperature = no *net* heat flow

2 Kelvin

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

3 Heating

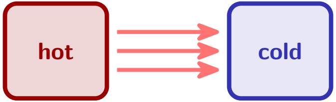
$Q = mc\Delta T$ – temperature changes

4 Changing state

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

Check yourself

1. Convert $27\,^{\circ}\mathrm{C}$ to kelvin.
2. Which equation for melting ice at $0\,^{\circ}\mathrm{C}$?
3. Why is a flat part of a heating curve flat?
4. Does the hotter body always hold more energy?



Answers **1.** $300.15\,\mathrm{K}$ **2.** $Q = mL_{\mathrm{f}}$ **3.** energy breaks bonds, not raising T **4.** no