Thermal Physics

Name _____ Date _____

By the end of this topic I can...

- use the Celsius and Kelvin scales, and describe the kinetic model of solids, liquids and gases
- use specific latent heat and specific heat capacity
- define internal energy and explain how it changes during heating and changes of state
- use the first law of thermodynamics $\Delta U = Q + W$ and $W = p\Delta V$
- describe thermal equilibrium, absolute zero and how thermometers work

Thermal basics

Temperature scales

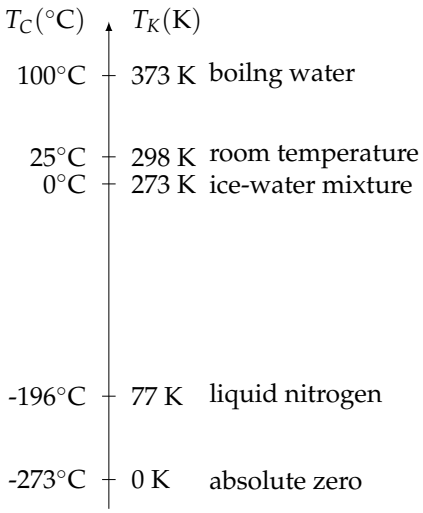
• **Celsius scale** (unit °C): 0 °C is the temperature of an ice–water mixture; 100 °C is the temperature of boiling water

• **Kelvin scale** (unit K): 0 K is absolute zero, the lowest temperature possible

• conversion:

$$T_K\,(\mathrm{K}) \underset{+273.15}{\overset{-273.15}{\rightleftharpoons}} T_C\,(^\circ\mathrm{C})$$

• a change of 1 °C equals a change of 1 K



Some temperatures on the two scales.

Kinetic theory of matter

The **kinetic theory of matter** explains *macroscopic* properties (density, compressibility, fluidity) from the *microscopic* behaviour of molecules.

• solid: molecules close together, tightly bonded, vibrate about fixed positions

• liquid: molecules quite close, vibrate but with some freedom to move about

• gas: molecules widely separated, free from neighbours, move rapidly

Specific latent heat

Changing state at constant temperature needs energy: the **latent heat**. It depends on the mass:

$$Q = Lm$$

Specific latent heat L is the thermal energy required to change the state of unit mass of a substance with no change in temperature.

- unit: $[L] =$ J kg^{-1}
- L depends only on the type of substance; for melting it is the specific latent heat of fusion, for boiling of vaporisation, and L_{vap} is greater than L_{fus}

Example 22.1 A 3.0 kW kettle holds 0.50 kg of water already at its boiling point. Neglecting heat losses, how long does it take to boil dry? ($L = 2.26 \times 10^6$ J kg^{-1})

$$Q = mL = 0.50 \times 2.26 \times 10^6 = 1.13 \times 10^6 \text{ J}$$
$$t = \frac{Q}{P} = \frac{1.13 \times 10^6}{3.0 \times 10^3} \approx 380 \text{ s} \approx 6.3 \text{ min}$$

Example 22.2 A student measures L for ice with an electric heater, but the insulation is imperfect. Two runs at different powers give: test 1: 60 W, 3.0 min, 40.4 g melted; test 2: 90 W, 3.0 min, 56.6 g melted. Find L.

Heat also enters from the surroundings: $mL = (P_{\text{heater}} + P_{\text{sur}})\, t$

$$40.4\, L = (60 + P_{\text{sur}}) \times 180$$
$$56.6\, L = (90 + P_{\text{sur}}) \times 180$$

Subtracting: $16.2\, L = 30 \times 180 \quad \Rightarrow \quad L \approx 333$ J g^{-1}, and $P_{\text{sur}} \approx 14.8$ W.

Question 22.1 In an experiment to find L for ice, ice at 0 °C is melted twice. Heater off: 10.0 min, 14.3 g melted. Heater on: 21 000 J supplied in 5.0 min, 70.0 g melted. (a) Why are two sets of readings taken? (b) Find L.

(a) The first run measures the melting caused by heat from the surroundings alone, so it can be subtracted from the second.
(b) In 5.0 min the surroundings melt $14.3/2 = 7.15$ g, so the heater melted $70.0 - 7.15 = 62.85$ g:

$$L = \frac{21\,000}{0.06285} \approx 3.3 \times 10^5 \text{ J kg}^{-1}$$

Specific heat capacity

Heating a substance (with no change of state) raises its temperature:

$$Q = cm\Delta T$$

_____Specific heat capacity_____ c is the thermal energy required _____per unit mass_____ to raise the temperature by _____one unit_____ (1 K).

• unit: $[c] =$ _____$\text{J kg}^{-1}\text{ K}^{-1}$_____; c does not depend on the size or shape of the sample

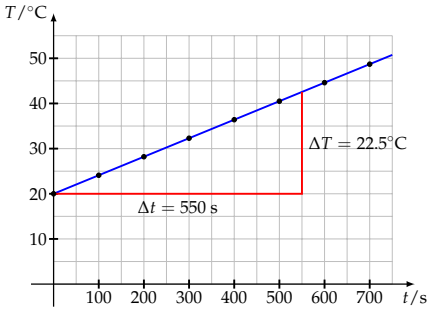
Example 22.3 30 g of ice at −20 °C is added to 270 g of water at 80 °C. No energy is lost. Find the final temperature. ($c_{\text{water}} = 4200$, $c_{\text{ice}} = 2100$ J kg^{-1} K^{-1}; $L_{\text{ice}} = 3.3 \times 10^5$ J kg^{-1})

Energy lost by hot water = energy gained by ice (warm to 0 °C, melt, warm to T):

$$4200 \times 0.27 \times (80 - T) = 2100 \times 0.030 \times 20 + 3.3 \times 10^5 \times 0.030 + 4200 \times 0.030 \times T$$

$$90\,720 - 1134\,T = 1260 + 9900 + 126\,T$$

$$T = \frac{90\,720 - 1260 - 9900}{1134 + 126} \approx 63\,^\circ\text{C}$$



Example 22.4 A 1.00 kg aluminium block is heated by a heater with current 4.2 A and p.d. 12 V. The graph shows the temperature rising. Find the specific heat capacity of aluminium.

$IV\Delta t = cm\Delta T \quad \Rightarrow \quad c = \dfrac{IV}{m\,\Delta T/\Delta t}$

gradient of the line: $\dfrac{\Delta T}{\Delta t} = \dfrac{22.5}{550} \approx 4.09 \times 10^{-2}\ ^\circ\text{C s}^{-1}$

$c = \dfrac{4.2 \times 12}{1.00 \times 4.09 \times 10^{-2}} \approx 1230\ \text{J kg}^{-1}\ ^\circ\text{C}^{-1}$

Question 22.2 A mixture is 5% silver and 95% gold by mass. Molten gold is mixed with silver at 20 °C so that the final mixture is at the melting point of gold (1340 K). Find the initial temperature of the gold. Silver: melting point 1240 K, $c = 235\ \text{J kg}^{-1}\ \text{K}^{-1}$, $L = 105\ \text{kJ kg}^{-1}$. Gold: $c = 129\ \text{J kg}^{-1}\ \text{K}^{-1}$ (solid or liquid).

Per kilogram of mixture, the silver (0.05 kg) must warm from 293 K to 1240 K, melt, then warm to 1340 K:

$$Q = 0.05\left[235 \times 947 + 1.05 \times 10^5 + 235 \times 100\right] \approx 1.76 \times 10^4\ \text{J}$$

This is supplied by 0.95 kg of gold cooling from T to 1340 K:

$$0.95 \times 129 \times (T - 1340) = 1.76 \times 10^4 \quad \Rightarrow \quad T \approx 1.48 \times 10^3\ \text{K}$$

Internal energy

Internal energy is the sum of the random kinetic energy of the molecules and the potential energy between the molecules:

$$U = E_k + E_p$$

It is a **state function**: it depends only on the current state, not on how it got there.

• **kinetic** part: from random motion only (not the motion of the whole system); the mean K.E. of molecules is

proportional to the temperature : $E_k \propto T$

• **potential** part: from forces *within* the system (not external fields); greater intermolecular separation means greater E_p, and breaking bonds also increases E_p

• E_p (gas) $>$ E_p (liquid) $>$ E_p (solid)[1]

[1] Ice melting is a counter-example for the separation: water is denser than ice. E_p still rises because hydrogen bonds are broken.

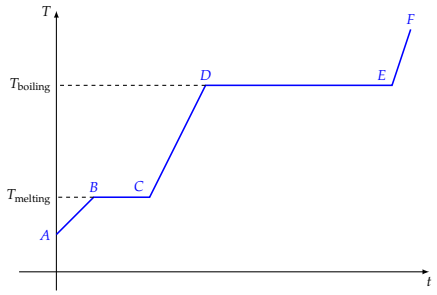
Internal energy of an ideal gas

No intermolecular forces, so $E_p =$ 0 and the energy is purely kinetic. Each molecule has $\frac{3}{2}kT$, so for N molecules

$$U = \tfrac{3}{2}NkT$$

the internal energy of an ideal gas is directly proportional to its temperature.

Change of state



Temperature against time for steady heating.

• AB (solid heating): T rises so E_k increases; separation unchanged so E_p is constant

• BC (melting): T constant so E_k constant; bonds break so E_p increases

• CD (liquid heating): E_k increases, E_p constant

• DE (boiling): E_k constant, E_p increases

• EF (gas heating): E_k increases; no forces so $E_p = 0$ constant

Question 22.3 Why is the specific latent heat of vaporisation much greater than that of fusion?

In vaporisation the molecules are separated completely (all bonds broken, and the substance expands against the atmosphere); in fusion the bonds are only loosened and the separation hardly changes. So a much larger increase in potential energy is needed.

Evaporation

Evaporation: a liquid changes to gas without boiling. Molecules move at various speeds; the fastest near the surface escape.

• **cooling effect**: the most energetic molecules leave, so the mean E_k of those remaining falls and the liquid cools

• the rate increases with temperature and with surface area

	boiling	evaporation
where	throughout the liquid	at the surface only
temperature	at the boiling point	at any temperature
bubbles	formed	none
rate	fast	slow

First law of thermodynamics

The first law: the increase in internal energy of a system equals the heat supplied to it plus the work done on it: $\Delta U = Q + W$

It is the law of conservation of energy. Sign conventions:

- $Q > 0$ if heat is supplied to the system; $Q < 0$ if it is released
- $W > 0$ if work is done *on* the system (compression, V decreases); $W < 0$ if the system does work (expansion, V increases)

Heat: $Q = cm\Delta T$ with no change of state, $Q = Lm$ during one. Work at constant pressure:[2]

$$W = p\,\Delta V \quad \text{(magnitude)}$$

If the volume does not change, no work is done.

Example 22.5 A gas is heated with 25 kJ of energy and expands by 0.10 m^3 at a constant pressure of 150 kPa. Find the change in internal energy.

Work: $|W| = p\Delta V = 150 \times 0.10 = 15$ kJ; the gas expands, so $W = -15$ kJ.
$\Delta U = Q + W = (+25) + (-15) = +10$ kJ

[2] If p varies, $W = \int p\,dV$: the area under the p–V graph.

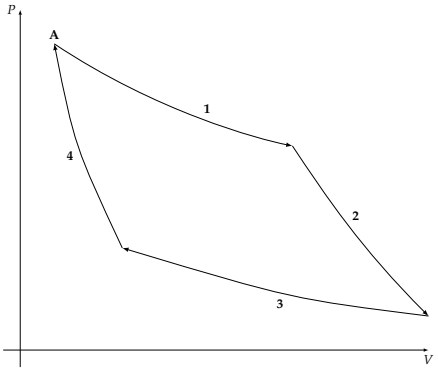
Example 22.6 Use internal energy and the first law to explain why boiling water needs a heat supply.

Boiling is at constant temperature, so $\Delta E_k = 0$; the molecules separate, so $\Delta E_p > 0$. Hence $\Delta U > 0$.
The volume increases, so work is done against the surroundings: $W < 0$.
$\Delta U = Q + W$ with $\Delta U > 0$ and $W < 0$ requires $Q > 0$: heat must be supplied.

Example 22.7 Explain, using the first law, why the air in a bicycle tyre gets warmer as you pump.

Compressing the gas: $W > 0$. Each stroke is quick, so little heat transfer: $Q \approx 0$.
$\Delta U = Q + W > 0$. For a gas $\Delta E_p = 0$, so $\Delta E_k > 0$.
Mean K.E. is proportional to temperature, so the temperature rises.

Gas cycles



A Carnot cycle, from A.

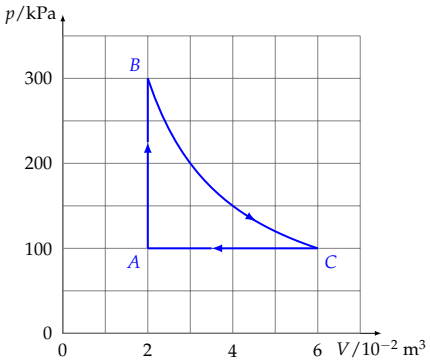
The cycle returns to A, so over the cycle $\Delta U =$ 0.

- 1 isothermal expansion : $\Delta U = 0$; gas does work so it must absorb heat
- 2 adiabatic expansion : $Q = 0$; gas does work so U and T fall
- 3 isothermal compression : $\Delta U = 0$; work done on the gas so it loses heat
- 4 adiabatic compression : $Q = 0$; work done on the gas so U and T rise

Example 22.8 Complete the table for the cycle: stage 1: Q =(a), $W = -936$ J, ΔU =(b); stage 2: $Q = 0$, W =(c), ΔU =(d); stage 3: Q =(e), $W = +702$ J, $\Delta U = 0$; stage 4: Q =(f), $W = +844$ J, $\Delta U = +844$ J.

Stage 1: isothermal, so (b) $= 0$ and (a) $= +936$ J.
Whole cycle: $\Delta U = 0$, so (d) $= -844$ J. In stage 2, $Q = 0$ so $\Delta U = W$, giving (c) $= -844$ J.
Stage 3: (e) $= -702$ J. Stage 4: no heating, (f) $= 0$.

Example 22.9 0.80 mol of an ideal gas goes round the cycle $ABCA$ shown. $T_A = 300$ K and the magnitude of the work from B to C is 6570 J. For each stage find W, Q and ΔU.



Cycle $ABCA$ for 0.80 mol of ideal gas.

Work: $A \to B$: no volume change, $W = 0$. $B \to C$: expansion, $W = -6570$ J. $C \to A$: compression at 100 kPa, $W = +1 \times 10^5 \times 4 \times 10^{-2} = +4000$ J.
ΔU (ideal gas, $\frac{3}{2}Nk\Delta T$): $A \to B$: same V, p triples so $T_B = 900$ K and $\Delta U = \frac{3}{2}nR\Delta T = \frac{3}{2} \times 0.80 \times 8.31 \times 600 \approx +5980$ J. $B \to C$: pV unchanged so $\Delta T = 0$, $\Delta U = 0$. $C \to A$: $\Delta U = -5980$ J.
First law $Q = \Delta U - W$: $Q_{AB} = +5980$ J, $Q_{BC} = +6570$ J, $Q_{CA} = -5980 - 4000 = -9980$ J.

Question 22.4 Show that the thermal energy needed to raise the temperature of n mol of an ideal monatomic gas by 1.0 K at constant volume is $\frac{3}{2}nR$.

No work is done at constant volume, so $Q = \Delta U = \frac{3}{2}Nk\Delta T = \frac{3}{2}nR\Delta T$. For $\Delta T = 1.0$ K, $Q = \frac{3}{2}nR$.

Question 22.5 Two identical balloons A and B hold the same gas at the same temperature and receive the same heat. A's volume is fixed; B may expand. Compare their final temperatures.

A: $W = 0$, so $\Delta U_A = Q$. B expands and does work: $W < 0$, so $\Delta U_B = Q - |W| < \Delta U_A$. Since $\Delta U \propto \Delta T$ for an ideal gas, $T_A > T_B$.

Temperature and thermometers

Thermal equilibrium

Heat flows spontaneously from hot to cold. Two objects at the same temperature are in **thermal equilibrium**: there is no net heat transfer.

Zeroth law: if A and B are each in thermal equilibrium with C, then A and B are in thermal equilibrium with each other.

Question 22.6 A student says temperature measures the amount of heat in an object. Why is this wrong?

Temperature measures the mean random kinetic energy per molecule, and shows the direction of heat flow. The thermal energy in an object depends also on its mass and substance: a bath of water at 40 °C holds far more internal energy than a hot spark at 1000 °C.

Absolute zero

Mean molecular K.E. is a microscopic measure of T, so the lowest temperature occurs when the molecules have minimum kinetic energy: **absolute zero**.[3]

The Kelvin scale is absolute: its zero does not depend on any particular substance. More precisely:

$$T_K\,(\mathrm{K}) \xrightarrow[+273.15]{-273.15} T_C\,(^\circ\mathrm{C})$$

(273.15 here; the rest of the course uses 273.)

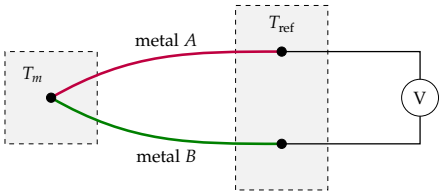
[3] Quantum effects mean the K.E. is never exactly zero, and 0 K cannot be reached in a finite number of steps (third law).

Thermometers

• **liquid-in-glass**: a liquid expands as it warms

• **resistance temperature detector**: resistance of a metal wire or thermistor changes with temperature

• **thermocouple**: a thermoelectric voltage appears between junctions at different temperatures; one junction is at a known reference temperature



A thermocouple.

Features of a thermometer: *range*, *sensitivity* (small changes detected), *response time*, *linearity*.

	liquid-in-glass	RTD wire	thermistor	thermocouple
range	narrow	wide	narrow	very wide
sensitivity	low	fair	high	high
response time	slow	fast	fast	very fast
linearity	good	good	limited	non-linear