# THERMAL PHYSICS

Thermodynamics - macroscopic behaviour of a system

Kinetic Theory - microscopic particular behaviour

## Absolute Temperature Scale

The lowest possible theoretical temperature is called ABSOLUTE ZERO
This is defined as -273.15 C
At 0 K (Kelvin) all molecules have NO KINETIC ENERGY.
1 degree Kelvin is equivalent to 1 degree Celsius in magnitude.

## Internal Energy

Fundamental property of thermodynamics.
Internal Energy, U, is the SUM of the randomly distributed KINETIC and POTENTIAL energies within a body.

$$U = \boxed{\Sigma E_k} + \boxed{\Sigma E_p}$$

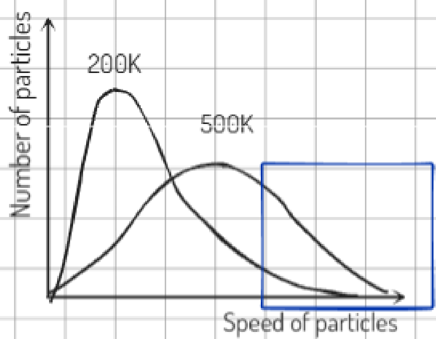
associated with the movement of particles within the body

associated with all fundamental force interactions

Ek links to the TEMPERATURE of the body.
Ep links to intermolecular forces and particle positions.

As the temperature of an object increases, the AVERAGE speed of the particles increases.
So the AVERAGE kinetic energy of the particles increases.

At higher temperatures the distribution becomes more even.

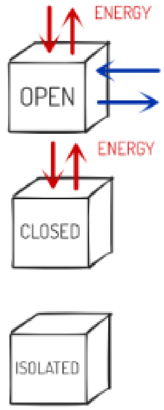



Note: during process such as sweating, the highest speed particles escape, reducing the mean kinetic energy

# Systems

A system is a group of bodies considered as a whole.

There are three types of THERMODYNAMIC system.
- OPEN SYSTEMS allow exchange of matter and energy from the surroundings.
- CLOSED SYSTEMS can exchange energy with the surroundings, but not matter.
- ISOLATED SYSTEMS exchange no matter or energy with its surroundings.




The total energy within a closed system is constant without external heating or cooling.

## The Zeroth Law of Thermodynamics

Imagine three thermodynamic systems A, B and C.

If A is independently in THERMAL EQUILIBRIUM with B and C, then B is also in THERMAL EQUILIBRIUM with C.

Thermal equilibrium means that the rate of energy transfer to and from the system (or between systems) are the SAME.

## The First Law of the Thermodynamics

When HEAT is supplied to a system it can lead to:

- A change in the internal energy of the system, with particles gaining kinetic or potential energy. This can lead to an increase in temperature or change in state.

- The system expanding and doing work on its surroundings.

The HEAT supplied to a system is equal to the SUM of the change in INTERNAL ENERGY of a system and the WORK DONE BY THE SYSTEM on its surroundings.

$$\boxed{Q} = \boxed{\Delta U} + \boxed{W}$$

| Worked example: |
| --- |
| A system consist of a gas. It is compressed so that 100J of work are done on the gas. 50J of heat are transferred to the gas.<br>Calculate the change in internal energy of the gas (ΔU) |

$$Q = \Delta U + W$$
$$\Delta U = Q - W$$
$$\Delta U = 50 - (-100)$$
$$\Delta U = +150\ J$$

negative as work is being done ON THE GAS, not by the gas

## HEAT CAPACITY

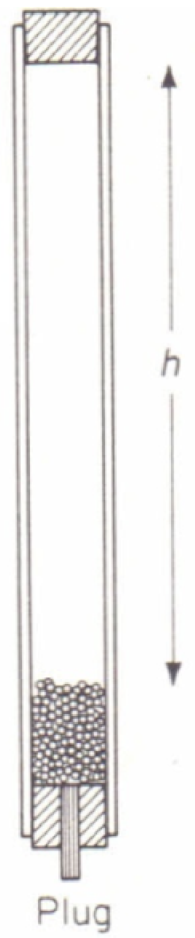
It takes different amounts of heat to increase the temperature of a unit mass of different substances.

The SPECIFIC HEAT CAPACITY of a substance is defined as the heat required to increase the temperature of a UNIT MASS of a substance by 1 K.

$$Q = mc\Delta T$$

Where:
Q = heat (J)
m = mass (kg)
ΔT = change in temperature (K)
c = specific heat capacity ($Jkg^{-1}K^{-1}$)

A positive Q represents heat being absorbed by a substance and an increase in temperature, whilst negative Q represents heat released and a decrease in temperature.




Worked example

A student tips some lead shot up and down in a sealed tube. After 20 turns, the temperature of the lead had risen by 1.5 K. Estimate the specific heat capacity of lead.

h = 1 m

loss of gpe from falling lead it transferred to shot as heat

$$mgh = mc\Delta T$$

$$gh = c\Delta T$$

$$c = \frac{gh}{\Delta T}$$

$$c = \frac{9.81 \times 20}{1.5} = 130\ Jkg^{-1}K^{-1}$$

Worked example

An immersion heater operates at 240 V and draws a current of 10 A. How long does it take to heat 120 kg of water from 293 K to 353 K?

$$IVt = mc\Delta T$$

$$t = \frac{mc\Delta T}{IV}$$

$$t = \frac{120 \times 4181 \times 60}{10 \times 240} = 12543\ s = 3.5\ hrs$$

Worked example

An aluminium sphere of mass 25g is at 500°C, and is plunged into 600g of water at 20°C. Calculate the final temperature of the water, stating any assumptions you make.

- Assuming all heat from the aluminium is transferred to the water and no energy is transferred to the $\boxed{\text{container}}$ or surroundings.
- The aluminium and water are in thermal equilibrium at the end.

$$m_{Al}c_{Al}\Delta T_{Al} + m_w c_w \Delta T_w = 0$$

(heat released by Al -ve) (heat gained by water +ve)

$$0.025 \times 897 \times (T - 500) + 0.600 \times 4181 \times (T - 20) = 0$$
$$22.4 \times (T - 500) + 2509 \times (T - 20) = 0$$
$$22.4T - 11200 + 2509T - 50180 = 0$$
$$2513.4\,T - 61380 = 0$$
$$T = 24.42\,^oC$$

Note: there may be questions where we need to include the container as well!

## SPECIFIC LATENT HEAT

When a substance changes state energy is either absorbed or released.
This depends on the mass of the substance.
THERE IS NO CHANGE IN TEMPERATURE DURING A CHANGE OF STATE.

The SPECIFIC LATENT HEAT is defined and as the energy required to change the state of a UNIT MASS of a substance.

$$Q = mL$$

Where:
Q = Heat (J)
m = mass (kg)
L = specific latent heat ($Jkg^{-1}$)

When a unit mass of a liquid FREEZES the atoms or molecules that form a solid structure (usually a lattice). Latent heat is RELEASED in this process and the INTERNAL ENERGY of the substance DECREASES.
We represent this as a NEGATIVE latent heat value.

This is known as the LATENT HEAT OF FUSION.
The latent heat of fusion for water is $340\ kJkg^{-1}$

When ice MELTS to form water it ABSORBS this much LATENT HEAT.
This is a POSITIVE value as internal energy is increasing.

When a liquid VAPORISES to form a gas, this is known as the LATENT HEAT OF VAPORISATION.
The latent heat of vaporisation for water is $2250\ kJkg^{-1}$

| How much energy is released when 100g of water boils? |
| --- |
| $Q = mL$<br>$Q = 0.1 \times (-2250 \times 10^3)$<br>$Q = -225 \times 10^3\ J$ |

100g of ice at $-5\ ^oC$ is mixed with 900g of water at $30\ ^oC$
Calculate the final temperature.

| heat to warm ice to melting point | + | heat to melt ice | + | heat to warm up melted ice | = | heat released by cooling water |
|---|---|---|---|---|---|---|

$$m_i c_i (0 - T_i) + m_i L_f + m_i c_w (T_f - 0) = m_w c_w (T_w - T_f)$$

$$(0.100 \times 2100 \times (0 - -5)) + (0.100 \times 340000) + (0.100 \times 4190 \times (T - 0)) = (0.900 \times 4190 \times (30 - T))$$

$$1050 + 34000 + 419T = 113130 - 3771T$$

$$4190T = 78080$$

$$T = 19\ ^oC$$