11

thermal properties of matter

content

Learning Outcomes:

Candidates should be able to:

(a) describe a rise in temperature of a body in terms of an increase in its internal energy (random thermal energy)

(b) define the terms *heat capacity* and *specific heat capacity*

(c) recall and apply the relationship *thermal energy* = *mass* × *specific heat capacity* × *change in temperature* to new situations or to solve related problems

(d) describe melting/solidification and boiling/condensation as processes of energy transfer without a change in temperature

(e) explain the difference between boiling and evaporation

(f) define the terms *latent heat* and *specific latent heat*

(g) recall and apply the relationship *thermal energy* = *mass* × *specific latent heat* to new situations or to solve related problems

(h) explain latent heat in terms of molecular behaviour

(i) sketch and interpret a cooling curve

11.1 Internal energy

Describe

Heat

- **Heat** (or *thermal energy*) may be transferred from one region (or body) to another as a result of a difference in *temperature*.

- When two bodies are placed in *thermal contact*, the warmer body becomes cooler while the cooler body becomes warmer, *i.e.*, **heat** flows spontaneously from the warmer to cooler regions.

Internal energy

- Every minute particle (such as ions, electrons, atoms or molecules) in a body has *potential* energy (E_p), due to their state and position, and, *kinetic* energy (E_k), due to their motion. Collectively, the sum of these energies is termed the **internal energy** (U) of the body, *i.e.*,

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

  Potentially, the *internal energy* of a system (a collection of minute particles) can be released for use. Usually it manifests the energy transfer in the form of *heat flow*.

- The *temperature* of a body is a measure of the *average kinetic energy* of the *minute particles* of the body. A rise in the *temperature* is therefore associated with an increase in the average *kinetic energy* of the particles, which must therefore imply an increase in the *internal energy* of the body.

Heat flow

- When a body is heated, its associated atoms or molecules start to move faster. *i.e.*, their *kinetic* energies are increased. So, in microscopic level, *heat* energy is stored in the form of *kinetic* energies in the atoms or molecules.

- The *kinetic energies* of the atoms or molecules is described as *random* **thermal energy**, to avoid confusion with the *kinetic* energy of the body as a whole.

- When the sum of *kinetic energies* in the particles increases, *thermal* energy and hence, *internal* energy increases, *temperature* being a gross measure of the state of the body increases.

- To increase the *internal energy* (mainly microscopic *kinetic energies*) of an object by a certain amount, the amount of *heat* to be supplied depends on its type of material, rise in temperature and mass, *i.e.*, capacity of absorbing or releasing *heat* varies from substance to substance.

☺ **Potential energy** of particles in a body tends to be small and does not change very much (unless there is a change of *state*, *e.g.*, liquid to gas); a change in **internal energy** is generally due to a change in **kinetic energy**.

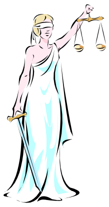
Worked Example

Example 1

Distinguish between internal energies of gases and solids microscopically.

Solution:

From a microscopic point of view, for a gas its internal energy may consist almost entirely of the kinetic energy of the gas molecules. It may also consist of the potential energy of these molecules in a gravitational, electric, or magnetic field. For a material solid, it may also consist of the potential energy of attraction or repulsion between the individual molecules of the material.



11.2 Specific heat capacity

Describe

Heat capacity

Some materials are harder to heat up than others. The molecules in a liquid such as water require <u>more</u> energy to move faster <u>than</u> do copper atoms in a solid. So, in order to record 1 °C increase in *temperature*, liquids would require more heat energy than solids.

> **Heat capacity**, ***C***, of a body is defined as the amount of thermal energy (Q) required to raise its temperature (θ) by one degree, without going through a change in state.

- SI unit of *heat capacity* is *joule per kelvin* or *joule per degree Celsius* [**J K^{-1}** or **J °C^{-1}**].

- In symbols, $\boxed{C = \dfrac{Q}{\Delta\theta}}$

where, C = heat capacity (J K^{-1} , J °C^{-1})

Q = heat or thermal energy absorbed or released (J)

$\Delta\theta$ = change in temperature (K or °C)

$(\theta_{\text{final}} - \theta_{\text{initial}})$ or $(T_{\text{final}} - T_{\text{initial}})$

Worked Example

Example 1

Calculate heat capacity of a piece of brass, if the heat of 8000 J is given to raise its temperature from 40 °C to 60 °C.

<u>Solution:</u>

Recall, heat capacity, $C = \dfrac{Q}{\Delta\theta}$

The values given: Q = 8000 J; $\Delta\theta$ = (60 – 40) = 20°C

Substituting: C = 8000/20 = 400 J°C^{-1}

Heat capacity of the brass piece is 400 J°C^{-1}. **(ans)**

Specific heat capacity

The greater the mass, the more atoms or molecules that need to be speeded up, and more *thermal energy* is needed to produce a given *temperature* rise. It is thus more common to consider the *heat capacity per unit mass* or *specific heat capacity* of the body.

Specific heat capacity, ***c***, of a body is defined as the amount of thermal energy (Q) required to raise the temperature (θ) of a unit mass of it by one degree, without going through a change in state.

- SI unit of *specific heat capacity* is *joule per kilogram per kelvin* or *joule per kilogram per degree Celsius* [**Jkg^{-1}K^{-1}** or **Jkg$^{-1\circ}$C^{-1}**].

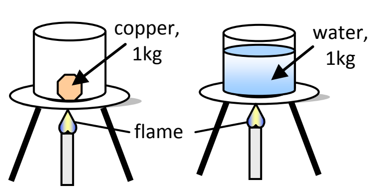
- In symbols, $\boxed{c = \dfrac{C}{m} = \dfrac{Q}{m \times \Delta\theta}}$ or $Q = m \times c \times \Delta\theta$

$$\begin{bmatrix}\text{thermal}\\\text{energy}\end{bmatrix} = [\text{mass}] \times \begin{bmatrix}\text{specific}\\\text{heat}\\\text{capacity}\end{bmatrix} \times \begin{bmatrix}\text{change in}\\\text{temperature}\end{bmatrix}$$

where, c = specific heat capacity (Jkg^{-1}K^{-1}, Jkg$^{-1\circ}$C^{-1})
C = heat capacity (JK^{-1} or J $^{\circ}$C^{-1})
m = mass of the substance (kg)
Q = heat or thermal energy absorbed or released (J)
$\Delta\theta$ = change in temperature (K or °C)

- *Specific heat capacity* of gases is higher than that of liquids and much higher than that of solids. The substances with higher *specific heat capacity* cool or warm very slowly compared to the substances with lower *specific heat capacity*.

- The *specific heat capacity* of water (liquid) is 4200 Jkg$^{-1\circ}$C^{-1} and for copper (solid) it is 400 Jkg$^{-1\circ}$C^{-1}.



copper takes less time than water to increase its temperature by say, 5 °C

☺ The **temperature** of an object gives a measure of how hot or cold it is, but it is not a measure of how much *internal energy* the object contains.

☺ With gases, the **molar heat capacity** (the *heat capacity* of one mole of a gas at constant pressure or constant volume) is generally more useful than the *specific heat capacity*, which is based on *mass*.

Worked Examples

Example 1

Calculate the required amount of energy to raise the temperature of 0.15 kg of water by 20 °C. The *specific heat capacity* of water is 4200 Jkg^{-1} °C^{-1}.

Solution:

Recall, $Q = m \times c \times \Delta\theta$

The values given: m = 0.15 kg, c = 4200 Jkg^{-1}°C^{-1}, $\Delta\theta$ = 20 °C.

Substituting: Q **=** 0.15 kg × 4200 Jkg^{-1}°C^{-1} × 20 °C = 12600 J

12,600 J of thermal energy is required. **(ans)**

☺ When the *temperature* increases, the change in *temperature* and *energy* is taken to be positive.

Example 2

Calculate the amount of energy released from a 0.2 kg copper, when it cools from 30°C to 25°C. The specific heat capacity of copper is 400 Jkg^{-1}°C^{-1}.

Solution:

Recall, $Q = m \times c \times \Delta\theta$

The values given:

m = 0.2 kg, c = 400 Jkg^{-1}°C^{-1}, ϑ = (25-30) °C^{-1} = (–5) °C.

Substituting: Q = 0.20 kg × 400 Jkg^{-1}°C^{-1} × (–5)°C = –400 J

400 J of thermal energy is released. **(ans)**

☺ When the *temperature* decreases, change in *temperature* and *energy* are taken to be negative.

Example 3

What is the temperature when 0.030 kJ of energy is supplied to a 5.0 g of iron rod at 3.0°C.? The *specific heat capacity* of iron is 460 Jkg^{-1}°C^{-1}.

Solution:

Recall, $Q = m \times c \times \Delta\theta \quad \Rightarrow \Delta\theta = \dfrac{Q}{m \times c}$

Let final temperature be θ°C

$\Rightarrow$ $\therefore$ rise in temperature = $(\theta - 3.0)^{\circ}\mathrm{C}$

The values given: m = (5.0/1000) kg, c = 460 Jkg^{-1}°C^{-1}, Q = 30 J

Substituting: $(\theta - 3.0)^{\circ}\mathrm{C}$ = $\dfrac{30\,\mathrm{J}}{(5.0/1000)\,\mathrm{kg} \times 460\ \mathrm{Jkg^{-1}\,{}^{o}C^{-1}}}$

$$\theta - 3.0 = \frac{30 \times 1000}{5.0 \times 460}\,{}^{\circ}\mathrm{C} = 13\,{}^{\circ}\mathrm{C}$$

The current final temperature is 13°C **(ans)**

Practical importance of the high specific heat capacity of water

- The specific heat capacity (s.h.c.) of water is relatively <u>high</u> compared to alcohol (liquid) and metals (solids). Hence,

  - Water is suitable for use as **coolant** on cars.
  - The s.h.c. of the seas is higher than the s.h.c. of the land masses, it leads to **milder climate** in the coastal areas, *i.e.*, a favourite settlement choice for man, plants and animals alike.
  - **Warm-blooded animals** are possible as the large percentage of water in the body keeps the animal's temperature stable with little fluctuations.
  - Water is used as **fire extinguisher** due to its high heat capacity.