5 Heat

HEAT : A FORM OF ENERGY

Students have already studied in their earlier classes that heat is a form of energy which cannot be seen but can only be felt. During winter, we sit in the sun (or near a fire place) to feel warm. In each case, we cannot see heat coming from the sun (or fire) but we can experience it very easily.

FLOW OF HEAT

Heat always flows from a body at a higher temperature to a body at a lower temperature.

1. If hot water is mixed with cold water, heat flows from hot water to cold water.
2. Consider two bodies A and B. Let the temperature of body A be 80°C and that of body B be 30°C. On bringing the bodies A and B in contact, heat will flow from hot body A to the cold body B.

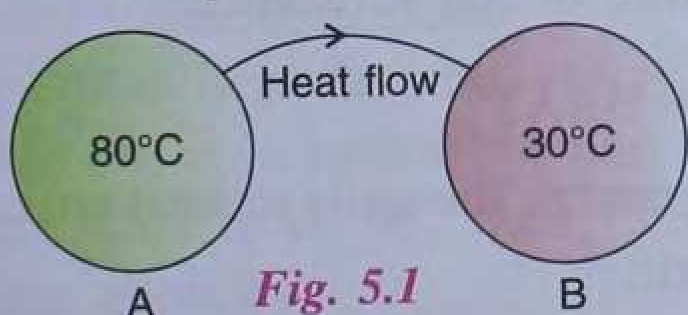


Fig. 5.1

The flow of heat from body A to body B will continue till both the bodies attain the same temperature.

It must be noted that when a body absorbs heat, its temperature rises and when it loses heat, its temperature falls. As shown in Fig. 5.1, body A loses heat, so its temperature falls whereas body B absorbs heat, so its temperature rises.

MEASUREMENT OF HEAT (Q)

The amount of heat gained or lost by a body depends upon the following factors :

1. *The amount of heat gained or lost by a body is directly proportional to its mass.*

$$Q \propto m. \qquad \text{.........(i)}$$

Experiment

Take 100 g and 200 g of water in two identical beakers. Keep each beaker on the wire gauze placed on tripod stands as shown in

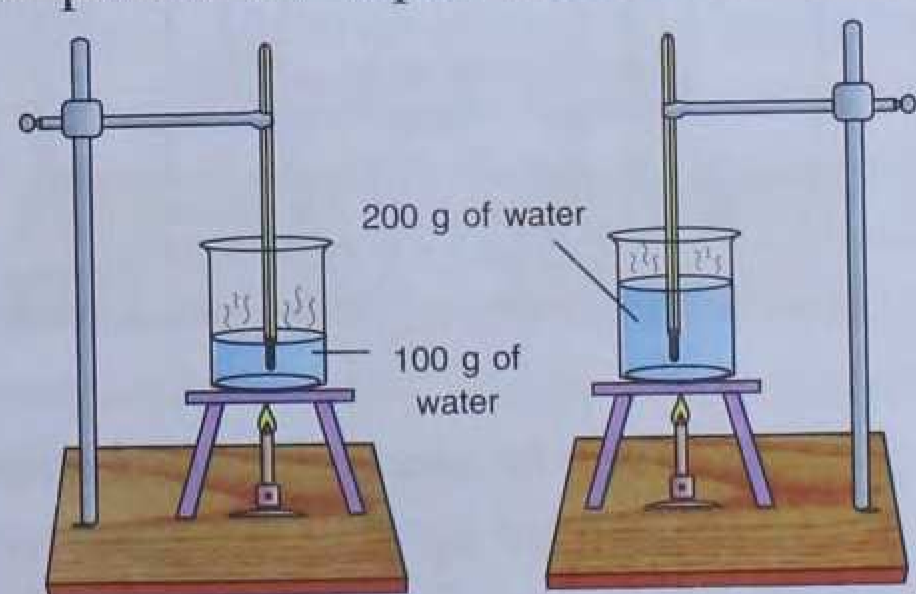


Fig. 5.2

Fig. 5.2. Put a thermometer in each beaker and heat the beakers on identical sources of heat for the same duration of time. You will notice that the rise in temperature of 100 g of water is nearly twice the rise in temperature of 200 g of water.

2. *The amount of heat gained or lost by a body depends upon the change in its temperature.*

$$Q \propto \Delta t$$

Take the same amount of water in two beakers of the same size, shape, *etc*. Place them on two tripod stands with a wire gauze on each of them. Put a thermometer in each beaker. Heat one beaker over a bigger source of heat and the other on a smaller one. You will notice that the rise in temperature of water in the first beaker will be more than that in the second beaker.

Since the first beaker is kept on the bigger flame (greater source of heat), it absorbs greater amount of heat (Q) and so it shows greater rise in temperature. Of course, both the beakers are heated for the same time.

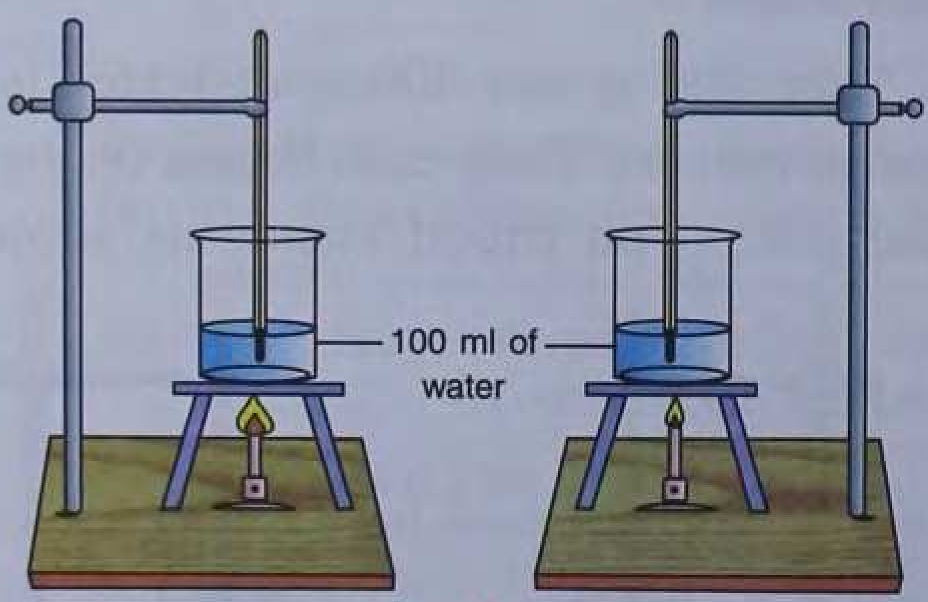


Fig. 5.3

3. *The quantity of heat absorbed depends upon the nature of the substance of the body.*

Take two identical beakers and put 50 g of water in one beaker and 50 g of mustard oil in the other. Put them on tripod stands with wire gauze. Put a thermometer in each and heat them on identical sources of heat for the same duration. You will notice that the rise in temperature in the two thermometers is different.

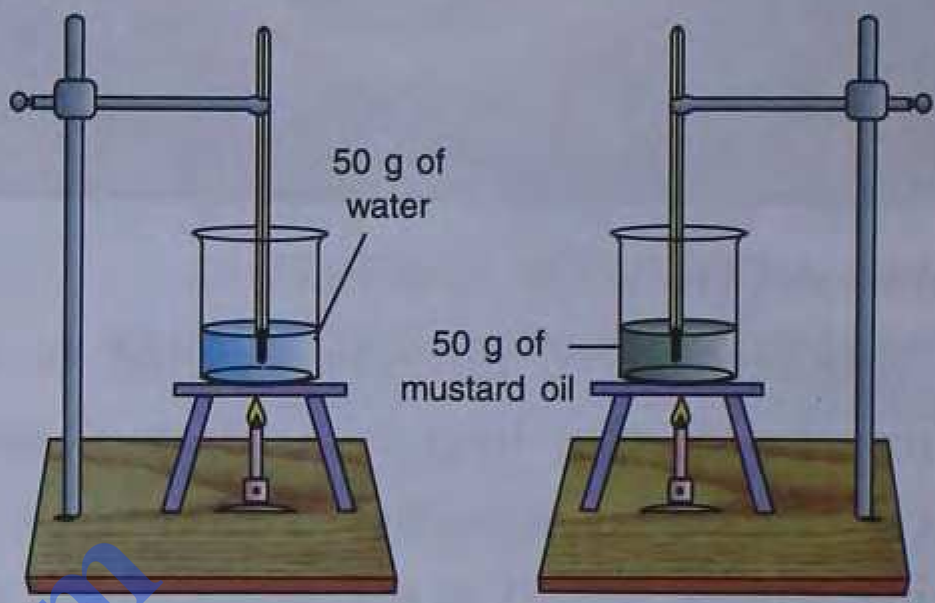


Fig. 5.4

Therefore, we conclude that quantity of heat (Q) gained (absorbed) by a body depends upon :

1. Mass of its substance (m)
2. Rise in temperature of the body (Δt)
3. Nature of the substance.

In the same way, when a body is cooled, it loses heat. Therefore, heat lost by a body depends upon :

1. Mass of its substance (m)
2. Fall in temperature of the body (Δt)
3. Nature of the substance.

UNIT OF HEAT

In general, the units of heat are calorie and kilocalorie.

One calorie is the amount of heat required to raise the temperature of 1 g of water by 1°C, *i.e.*, from 14·5°C to 15·5°C.

One kilocalorie is the amount of heat required to raise the temperature of 1 kg of water by 1°C *i.e.,* from 14·5°C to 15·5°C.

1 Kilocalorie = 1000 Calories

Since, heat is a form of energy, it can also be expressed in joule. Joule is the SI unit of energy. The relation between joule and other units of heat is given below :

1 Calorie = 4·2 Joule,

1 Kilocalorie = 4200 Joule and

1 Joule = 0·24 Calorie

SPECIFIC HEAT CAPACITY

If you heat the same amount of water and sand by similar sources and for the same duration, you will notice that the temperature of sand goes up more as compared to water. This is due to the capacity of different substances to absorb different amounts of heat *i.e.,* specific heat.

The quantity of heat required to raise the temperature of unit mass of a substance by 1°C is called its **specific heat capacity**. It is generally denoted by C.

∴ Specific heat capacity of the substance of a body

$$= \frac{\text{Quantity of heat absorbed by the body}}{\text{Mass of the body} \times \text{Rise in its temperature}}$$

$$\Rightarrow \quad \boxed{C = \frac{Q}{m \times \Delta t}}$$

If the unit of heat absorbed (Q) is in Joule, unit of mass (m) is in kg and the unit of rise in temperature (Δt) is in °C, then, the unit of specific heat capacity

$= \text{J/kg°C} = \text{J kg}^{-1}\ \text{°C}^{-1}$

S.I. unit of specific heat capacity

$= \text{J/kg K} = \text{J kg}^{-1}\ \text{K}^{-1}$

The specific heat capacity of water is 4200 J/kg°C. It means that 4200 J of heat is required to raise the temperature of 1 kg of water by 1°C.

The smaller unit of specific heat capacity is J/g°C = $\text{J g}^{-1}\ \text{°C}^{-1}$. *i.e.,* if the amount of heat absorbed is in joule, mass is in g (gramme) and rise in temperature is in °C; then the specific heat capacity is in $\text{J g}^{-1}\ \text{°C}^{-1}$.

Specific heat capacity of water is $4{\cdot}2\ \text{J g}^{-1}\ \text{°C}^{-1}$. It means that 4·2 J of heat energy is absorbed by 1 g water to raise its temperature by 1°C.

Specific heat capacities of some substances are given below :

1. Kerosene oil = $2100\ \text{J kg}^{-1}\ \text{°C}^{-1} = 2{\cdot}1\ \text{J g}^{-1}\ \text{°C}^{-1}$
2. Water = $4200\ \text{J kg}^{-1}\ \text{°C}^{-1} = 4{\cdot}2\ \text{J g}^{-1}\ \text{°C}^{-1}$
3. Mercury = $139\ \text{J kg}^{-1}\ \text{°C}^{-1} = 0{\cdot}139\ \text{J g}^{-1}\ \text{°C}^{-1}$
4. Iron = $483\ \text{J kg}^{-1}\ \text{°C}^{-1} = 0{\cdot}483\ \text{J g}^{-1}\ \text{°C}^{-1}$

Out of many substances, given above, water has exceptionally high specific heat capacity. It means that a large amount of heat is required to raise temperature of a given mass of water. In other words, it can be stated that water has greater value of specific heat capacity. There are many applications of high specific heat capacity of water in our daily life.

ADVANTAGES OF HIGH SPECIFIC HEAT CAPACITY OF WATER IN EVERYDAY LIFE

1. Formation of land-breeze and sea-breeze : The specific heat capacity of water is about five times more than that of sand. During the day, when sun rays fall on sand and water, the sand gets heated more quickly as compared to water. The air layers associated with sand get heated and rise up and the sea-breeze moves towards land to fill up the gap. At night, sand cools much faster than water.

The air layers above water are hotter than the air layers above sand. Therefore, they rise up and the land breeze moves towards the sea to fill up the gap.

Hence, a moderate climate is maintained near the sea-shore.

> *Note :* When temperature of a body rises, it abosorbs a certain amount of heat. Conversely, when temperature of the same body falls equally, it gives out the same amount of heat.

Do You Know ?

> Moderate climate in coastal areas is the result of slow absorption of solar heat energy by water during the day and slow release of heat energy during the night.
>
> The vast oceans on earth (about 75% of the surface area) are responsible for tempering the climate on earth permitting life to exist.

2. **Fomentation :** As water has high specific heat capacity, it absorbs a large amount of heat when its temperature rises. Conversely, it gives out large amount of heat when its temperature decreases. Also, due to high specific heat capacity, it takes a longer time to cool down. Due to this reason we use water in hot bags for fomentation purposes, as it keeps us warm for a longer period of time.

3. **As Coolant :** Water is used as an effective coolant by flowing it through pipes around the heated parts of a machine, as in the radiator of a car. This is because for the same rise in temperature, water will extract more amount of heat in comparision to other liquids owing to its high specific heat capacity.

4. **To Prevent Freezing :** In cold countries, wine and juice bottles are kept under water to avoid freezing because water does not cool quickly owing to its high specific heat capacity.

5. Farmers protect their crops from frost by filling the fields with water. Since the specific heat capacity of water is high, the temperature of water does not decrease much on a cold night. Thus, it does not allow the temperature of the surrounding area of the plants to decrease.

6. Metals like copper, aluminium, iron, *etc.*, are good conductors of heat and they have low specific heat capacity. Therefore, the cooking utensils are made up of these metals so that they can quickly attain a high temperature for a given quantity of heat supplied.

We know, specific heat capacity of the substance of a body is expressed as

$$\boxed{C = \frac{Q}{m \times \Delta t}}$$

$$\Rightarrow \quad Q = m \times C \times \Delta t$$

where Q = heat given or taken by the body

m = mass of the body

C = specific heat capacity of the substance of the body

and Δt = change (rise or fall) in temperature

Example 1 :

What is the amount of heat required to raise the temperature of 5 kg of iron from 30°C to 130°C ? Specific heat capacity of iron = 483 J kg^{-1} °C^{-1}.

Solution :

$$\text{Given : } m = 5\ kg,\ C = 483\ \text{J kg}^{-1}\ {}^{\circ}\text{C}^{-1}$$

$$\text{Rise in temp. } \Delta t = 130^{\circ}C - 30^{\circ}C = 100^{\circ}\text{C}.$$

Heat gained, $Q = m \times C \times \Delta t$

$$= 5\ kg \times 483\ \text{J kg}^{-1}\ {}^{\circ}\text{C}^{-1} \times 100^{\circ}C$$

$$= \mathbf{241500\ J}$$

Example 2 :

A lead block of mass 2 kg is supplied with a heat of 3600 J. Calculate the rise in temperature, if the specific heat capacity of lead is 130 J $\text{kg}^{-1}\ {}^{\circ}\text{C}^{-1}$.

Solution :

$$\text{Given : } m = 2\ kg,\ C = 130\ \text{J kg}^{-1}\ {}^{\circ}\text{C}^{-1}$$

$$Q = 3600\ J$$

$$\text{We know, } Q = m \times C \times \Delta t$$

$$\therefore \text{ Rise in temp. } \Delta t = \frac{Q}{m \times C}$$

$$= \frac{3600\ \text{J}}{2\ \text{kg} \times 130\ \text{J kg}^{-1}\ {}^{\circ}\text{C}^{-1}}$$

$$= \frac{360\cancel{0}}{26\cancel{0}}\ {}^{\circ}\text{C} = \mathbf{13{\cdot}85^{\circ}C}$$

Example 3 :

400J of heat is required to raise the temperature of a certain mass of iron from 30°C to 35°C. Calculate the mass of iron, if specific heat capacity of iron is 480 J $\text{kg}^{-1}\ {}^{\circ}\text{C}^{-1}$.

Solution :

$$\text{Given : } Q = 400\ \text{J},\ C = 480\ \text{J kg}^{-1}\ {}^{\circ}\text{C}^{-1}$$

$$\Delta t = (35 - 30)^{\circ}\text{C} = 5^{\circ}\text{C}$$

$$\text{We know, } Q = m \times C \times \Delta t$$

$$\Rightarrow \quad m = \frac{Q}{C \times \Delta t}$$

$$= \frac{400\ \text{J}}{480\ \text{J kg}^{-1}\ {}^{\circ}\text{C}^{-1} \times 5^{\circ}\text{C}}$$

$$= \frac{400}{2400}\ \text{kg} = \mathbf{0{\cdot}167\ kg}$$

Example 4 :

How much heat is required to raise the temperature of 100 g of water from 5°C to 95°C (specific heat capacity of water = 1 cal/g°C).

Solution :

$$\text{Here } Q = ?,\ m = 100\ \text{g},\ \Delta t = 95^{\circ}\text{C} - 5^{\circ}\text{C} = 90^{\circ}\text{C}$$

$$C = 1\ \text{cal/g}^{\circ}\text{C}$$

$$Q = m \times C \times \Delta t$$

$$= 100\ \cancel{g} \times 1 \frac{\text{cal}}{\cancel{g}\cancel{^{\circ}C}} \times 90\cancel{^{\circ}C}$$

$$= \mathbf{9000\ calories}$$

Example 5 :

If 60 calories of heat is supplied to 15 g of water, what is the rise in temperature. Given, specific heat capacity of water = 1 cal/g °C.

Solution :

$$Q = 60\ \text{calories},\ m = 15\ \text{g},\ C = 1\ \text{cal/g}\ {}^{\circ}\text{C},\ \Delta t = ?$$

$$Q = m \times C \times \Delta t$$

$$\Rightarrow \quad \Delta t = \frac{Q}{m \times C} = \frac{60\ \cancel{cal}}{15\cancel{g} \times 1 \frac{\cancel{cal}}{\cancel{g}^{\circ}C}}$$

$$= \frac{60}{15}\ {}^{\circ}\text{C} = \mathbf{4^{\circ}C}$$

HEAT CAPACITY

The quantity of heat required to raise the temperature of a body through 1°C is called the **heat capacity** or the **thermal capacity** of the body.

$$\text{Heat capacity} = \frac{\text{Quantity of heat supplied}}{\text{Rise in temperature}}$$

The unit of heat (thermal) capacity is J/°C or Cal/°C. In the International System of units (SI) heat capacity is expressed as Joules per Kelvin, *i.e.*, J/K (or J K^{-1}).

Heat capacity of a body is 360 J/°C (or 360 J$^{\circ}$C^{-1}). It means that 360 J of heat is required to raise the temperature of the body by 1°C.

Similarly, heat (thermal) capacity of a body is 1·6 cal/°C (cal $°C^{-1}$). It means, 1·6 cal of heat is required to raise the temperature of the body by 1°C.

Heat (thermal) capacity of a body

$$= m \times C$$

where, m = mass of the body

and, C = specific heat capacity of the body

> From the above, we can say, heat capacity of a body is the amount of heat required to raise its temperature by 1°C and specific heat capacity of a substance is the amount of heat required to raise the temperature of unit mass of the substance of a body by 1°C. It is denoted by C.
>
> ***Note :*** Specific heat capacity is the property of the substance of a body whereas heat capacity is the property of the body.

For a body made of a given substance, heat capacity is directly proportional to the mass of the substance it contains.

— Doubling the mass of the substance in a body doubles its heat capacity.

PRINCIPLE OF CALORIMETRY

When two bodies (substances) at different temperatures are kept in contact, heat flows from the body at a higher temperature to the body at a lower temperature till the temperature of both the bodies is same. In this case, the heat given by the hot body is equal to heat taken by the cold body provided no heat is lost or gained from the surroundings, *i.e.*,

> Heat lost by hot body = Heat gained by cold body

Example 6 :

600 g copper at 50°C is mixed with 1000g water at 20°C. Find the final temperature of the mixture. Specific heat capacity of copper is 0·4 J g^{-1} $°C^{-1}$ and that of water is 4·2 J g^{-1} $°C^{-1}$.

Solution :

Let the final temperature of the mixture be x°C. Clearly, x°C is between 50°C and 20°C.

For Copper :

m = 600 g, C = 0·4 J g^{-1} $°C^{-1}$ and

fall in temperature, $\Delta t = (50 - x)$°C.

$$\text{Heat lost by copper} = m \times C \times \Delta t = 600 \times 0{\cdot}4 \times (50 - x)\text{J}$$

For Water :

m = 1000 g, C = 4·2 J g^{-1} $°C^{-1}$ and

rise in temperature, $\Delta t = (x - 20)$°C

$$\therefore \text{Heat gained by water} = m \times C \times \Delta t = 1000 \times 4{\cdot}2 \times (x - 20)\text{ J}$$

According to the principle of calorimetry :

$$\text{Heat lost by copper} = \text{Heat gained by water}$$

$$\Rightarrow 600 \times 0{\cdot}4 \times (50 - x) = 1000 \times 4{\cdot}2 \times (x - 20)$$

$$\Rightarrow 12000 - 240x = 4200x - 84000$$

$$\Rightarrow 4440x = 96000$$

$$\Rightarrow x = \frac{96000}{4440} = 21{\cdot}6$$

∴ Final temperature of the mixture = 21·6°C

CHANGE OF STATE

The change of a substance from one physical state to another is called its change of state. This can be achieved either by heating or cooling the substance. Figure 5.5 given below shows the complete cycle of change of state.

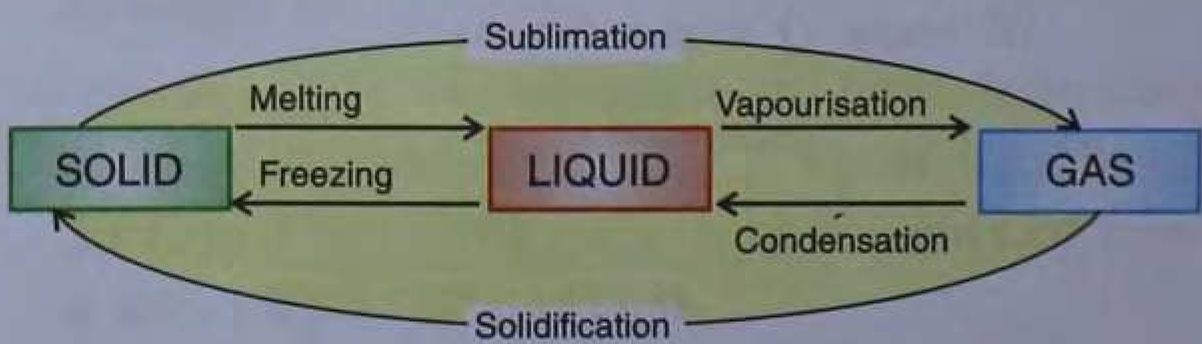


Fig. 5.5

FUSION OR MELTING

It is the process during which a solid changes to its liquid state at a constant temperature by absorption of heat. The fixed temperature at which melting occurs is known as the **melting point** of the solid.

For example :

Ice (solid) melts into water (liquid) at a temperature of 0°C.

VAPOURISATION OR BOILING

It is the process during which a liquid changes into its vapour state at a constant temperature by absorption of heat. This fixed temperature is known as the **boiling point** or **vapourisation point** of the liquid.

For example :

Water (liquid) vapourises to steam (gas) at a temperature of 100°C.

CONDENSATION

It is the process during which vapours change into liquid state at a constant temperature by releasing heat. This fixed temperature is known as the condensation temperature.

For example :

Steam (gas) condenses to water (liquid) at a temperature of 100°C.

FREEZING

It is the process during which a liquid changes into its solid state at a constant temperature by releasing heat. The fixed temperature is known as the **freezing point**.

For example :

Water (liquid) freezes into ice (solid) at a temperature of 0°C.

SUBLIMATION

It is the process in which a solid directly converts into its gaseous state at all temperatures.

For example :

Camphor or naphthalene directly turns into gaseous state at room temperature.

SOLIDIFICATION

It is the process in which a gas condenses directly into its solid state.

For example :

Carbon dioxide gas gets converted into dry ice.

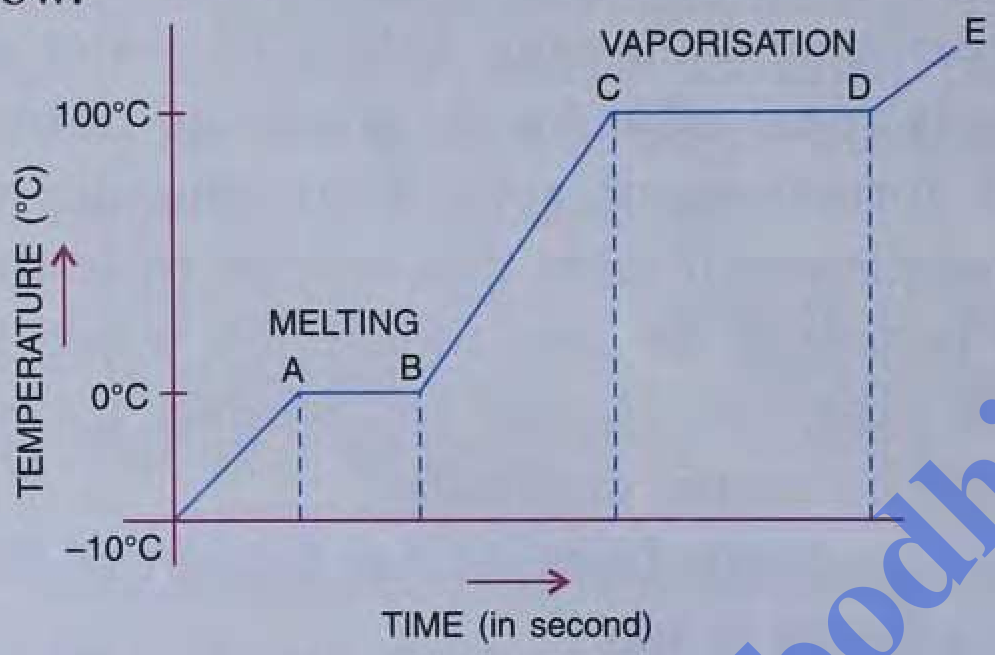
Note : It is very significant that ice expands and floats on water. The consequences of this action can be seen in broken water lines in the winter and pot holes in the roads. Freezing and thawing action of water are largely responsible for the fracturing of rocks and formation of soil. Also consider the consequences if lakes and streams froze from the bottom to the top, aquatic life would not even exist, and climate and weather patterns would be altered drastically.

Change of state from solid to liquid and then from liquid to gas :

Consider a certain amount of ice say at –10°C in a beaker. Place a thermometer inside the beaker. It reads –10°C initially. Now gradually heat the beaker.

We notice that the temperature shown by the thermometer increases from –10°C to 0°C and no ice melts during this period. When temperature is 0°C then the ice starts melting. The heat supplied is used by the ice to convert its state into liquid and the heat absorbed by the ice is called the **Latent Heat of Fusion**. Once all the ice converts into water at 0°C and we continue heating it, the temperature of water (water obtained on melting of ice) starts rising. The continuous supply of heat will raise the temperature of the water upto 100°C. On attaining the temperature of 100°C, the temperature of the water will stop rising. However, further supply of heat to this boiling water will start converting it into steam, at 100°C (*i.e.*, water will start converting from liquid to vapours). During this period, again the thermometer does not record any rise in temperature and it remains at 100°C. The heat absorbed during the change of state from liquid to gas is called **latent heat of vapourisation**.

This entire change in temperature with respect to time is shown in the graph shown below.



SPECIFIC LATENT HEAT

We know that during the process of fusion and vaporisation, the temperature remains constant although we continuously supply heat.

Where is this heat going away ?

This heat is getting utilised to convert the state of solid (or liquid) into liquid (or gas). The temperature will not rise until whole of the solid (or liquid) turns into liquid (or gas). Let us understand with the following example.

We take some ice in a beaker and put a thermometer into it to measure the temperature. It will be around 0°C (Fig. 5.6 a). Now we

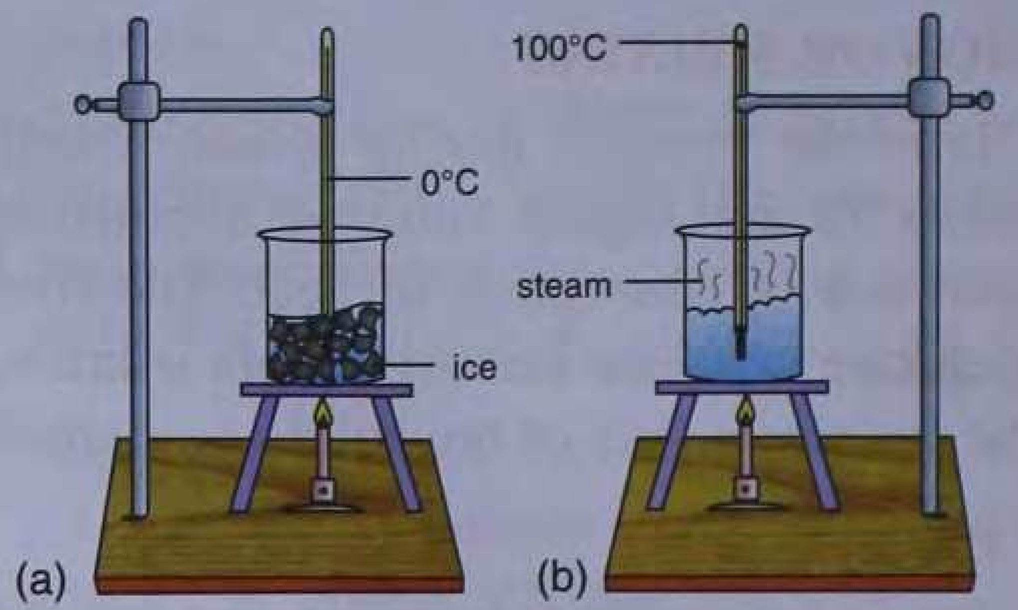


Fig. 5.6 (a) solid to liquid (fusion of ice) (b) liquid to gas (vaporisation of water)

heat it over a flame. On heating, ice starts melting but the thermometer reading does not change till whole of the ice melts. Similarly, when water starts boiling, the thermometer will show the reading around 100°C. Steam starts forming and we observe that the thermometer reading remains same till whole of the water converts into steam (Fig. 5.6 b).

The amount of heat required by a unit mass of a substance to change its state without any change in temperature is known as its **specific latent heat**.

Latent heat is the heat released or absorbed by a substance during its change of physical state without any change in temperature.

Note : The term 'latent heat' was introduced around 1750 by Joseph Black. It is derived from the Latin word latere (to lie hidden).

The use of ice to keep beverages cold in an insulated cooler is an example of the use of the latent heat of fusion of water.

— A tub of water in a green house on a cold winter night will moderate the temperature in the green house because of the latent heat released by the water in the process of freezing.

SPECIFIC LATENT HEAT OF FUSION — It is denoted by L_f

The specific latent heat of fusion of a substance is the quantity of heat required to convert a unit mass of the substance from solid to liquid state at its melting point without any change in its temperature.

SPECIFIC LATENT HEAT OF VAPOURISATION — It is denoted by L_v

The specific latent heat of vapourisation of a substance is the quantity of heat required to convert a unit mass of the substance from liquid state to vapour state at its boiling point without any change in its temperature.

The specific latent heat of fusion of a substance is 720 J kg^{-1}. It means that 720 J of heat is required to change the state of 1 kg substance, from solid to liquid, at a constant temperature (its melting point).

Similarly, specific latent heat of vapourisation of a substance is 1980 J g^{-1}. It means that 1980 J of heat is required to change the state (liquid to vapour) of 1 g substance at a constant temperature (its boiling point).

Note : The terms latent heat of fusion (melting) and latent heat of vapourisation (boiling) describes the direction of energy flow when changing from one phase to the next : from solid to liquid and to gas.

In both cases the change is endothermic, meaning that the system absorbs energy on going from solid to liquid to gas. The change is exothermic for the opposite direction.

Let (m) be the mass of a body and (L) be the specific latent heat of its substance. Then, heat (Q) required to change its state without any change in its temperature is given by,

$$Q = m \times L.$$

Latent heat is the amount of heat energy released or absorbed during the change of state of the body.

Consider a toy (body) made of 50 g of copper. Suppose we heat it strongly such that it acquires the temperature of 1085°C, the melting point of copper. The toy will become hot but on further heating it will start melting. On further heating, this process of melting will continue till the whole of the copper toy (the body) melts into its liquid state at 1085°C. **Remember during the process of melting the temperature of copper will remain the same, *i.e.*, 1085°C, which is the melting point of copper.**

In this case, the total amount of heat required to melt the whole of copper in the body (toy) is **latent heat of the body** whereas the heat required to melt a unit mass (1 g here) of copper is **specific latent heat of the copper**.

UNITS OF SPECIFIC LATENT HEAT

The SI unit of specific latent heat is J/kg or J kg^{-1} *i.e.*, if heat is in joule and mass of the body is in kg, the unit of specific latent heat is J kg^{-1}.

If heat is measured in joule and mass of the body is in g (gramme), the unit of specific latent heat is J/g or J g^{-1}.

Note : 1 J g^{-1} = 1000 J kg^{-1}.

Example 7 :

Calculate the amount of heat required to convert 200 g of ice at 0°C into water at 0°C. Specific latent heat of fusion of ice = 336 J g^{-1}.

Solution :

Given : m = 200 g and L = 336 J g^{-1}.

$\therefore$ **Amount of heat required, Q** = mL

= 200 × 336 J.

= **67200 J**

Example 8 :

2875 J of heat is required to melt 115 g of lead at its melting point. Calculate the specific latent heat capacity of fusion of lead.

Solution :

$$Q = m \times L$$

$$\Rightarrow \quad L = \frac{Q}{m} = \frac{2875}{115} = 25$$

$\therefore$ **Sp. latent heat of fusion of lead = 25 J g^{-1}**

Example 9 :

Find the amount of heat energy liberated when 20 g of steam at 100°C liquefies to form water at 100°C. Given, specific latent heat of vapourisation of steam is 2260 J g^{-1}.

Solution :

Mass of steam condensed (m) = 20g

Specific latent heat of vapourisation of steam = 2260 J g^{-1}

$\therefore$ **Heat liberated by steam (Q)** = mL

= 2260 × 20

= **45200 J**

Example 10 :

How much mass of water at 100°C must be heated so as to form steam at 100°C when a burner supplies 67800 J of heat energy ? Specific latent heat of vapourisation of water is 2260 Jg^{-1}.

Solution :

Mass of water (m) = ?

Specific latent heat of vapourisation of steam (L) = 2260 J g^{-1}

Heat supplied by burner (Q) = 67800 J

Since, heat supplied Q = m × L

$$\Rightarrow \quad 67800 = m \times 2260$$

$$\Rightarrow \quad m = \frac{67800}{2260} = 30 \text{ g}$$

NATURAL CONSEQUENCES OF HIGH SPECIFIC LATENT HEAT CAPACITY OF ICE

Ice has unusually high specific latent heat capacity of about 336000 J/kg (J kg^{-1}).

1. Due to high specific latent heat capacity of ice, snow on the mountains does not melt as a whole but melts gradually into water with the heat of the sun. If the specific latent heat of ice would not have been so high, all the snow would have melted very quickly and there would have been floods in the rivers.
2. Water in lakes and ponds in cold places do not freeze all at the same time. It freezes slowly and release of heat during freezing keeps the temperature of the surroundings moderate.
3. Drinks are cooled more effectively by ice pieces at 0°C and not by water at 0°C. This is because 1 g of ice takes away 336 J of heat from the drink to melt into water at 0°C.

Consequences of high specific latent heat capacity of steam :

1. Steam causes more severe burns than boiling water although both are at 100°C. The specific latent heat of steam is 2268 J/g. Therefore, 1 g of steam gives out 2268 J of extra heat energy to condense into water at 100°C, thereby causing more serious burns.
2. Steam is used for running trains or machines because the high amount of heat contained in it turns into mechanical energy.
3. Since water has exceptionally high specific latent heat of vaporisation, water from the soil does not evaporate quickly by heat of the sun. The plants are thus protected from wilting.

EVAPORATION AND VAPOURISATION

Both the processes convert liquid into gaseous state but there are four major differences between the two processes.

1. Evaporation takes place at all temperatures but vaporisation takes place only at a fixed temperature which is the boiling point of the liquid.
2. Evaporation is a slow and gradual process whereas vaporisation is a rapid and violent process.
3. Evaporation takes place only at the surface of the liquid whereas vapourisation takes place over the entire liquid.
4. Evaporation causes cooling but vapourisation does not produce cooling.

Factors affecting the rate of evaporation

The rate of evaporation depends on the following factors :

1. **Temperature of the liquid :** Rate of evaporation increases with the rise in temperature *e.g.* hot water evaporates faster than water at lower temperature.
2. **The nature of liquid :** Different liquids have different rates of evaporation *e.g.* petrol, ether, alcohol, etc. evaporate much faster than water.
3. **Surface area of liquid :** Rate of evaporation is faster if the surface area is more, *e.g.* hot milk on a saucer cools faster than when it is taken in a cap.
4. **Air layer flowing just above the liquid :** If the air layer just above the liquid is moving faster, the rate of evaporation of the liquid increases because of which the liquid cools faster. This is the reason that if we blow air above the milk surface, it cools faster.
5. **Temperature of surrounding :** Higher the temperature of the surrounding atmosphere greater will be the rate of evaporation. This is the reason why we spread clothes outside (especially in sun) to dry.
6. **Dryness of air :** If the air over the surface of liquid is dry, it evaporates faster. This is the reason why wet clothes dry faster in winter rather than in monsoon or summer.

Intext Questions

1. Why wet clothes have to be spread so as to dry quickly ?
2. How can you effectively show that evaporation causes cooling ?
3. Why do dogs hang out their tongue on a hot sunny day ?

Application of evaporation

1. A wet handkerchief is placed on the forehead of a person having high fever. The cloth takes the latent heat from forehead to evaporate the water thereby bringing the temperature down.
2. We often pour tea in a saucer to cool it faster. In the saucer the surface area increases and hence it cools faster.
3. Earthen pots are used in summer. The pot has several pores through which water rinses (ooze) out to the outer surface of the earthen pot (pitcher) and evaporates. In this process it takes away heat from the water inside the pitcher. As a result water in the pitcher gets cooled.

4. Evaporation of sweat from our body helps to maintain the body temperature at 37°C (98·4°F). When sweat evaporates, it requires heat which it takes away from our body thereby lowering the temperature.

The process during which a liquid changes into its gaseous state on heating is called **vapourisation**. If the temperature of a liquid increases, the pressure of molecules of the liquid also increases. When the pressure of molecules of liquid becomes equal to the atmospheric pressure, the molecules of liquid change into vapours.

The temperature at which pressure of molecules of a liquid becomes equal to the atmospheric pressure is called its **boiling point**.

CONDUCTORS AND INSULATORS

Substances through which heat can pass easily are called **good conductors** of heat. Metals are good conductors of heat, *e.g.*, silver, copper, aluminium, iron, etc. In fact, silver is the best conductor of heat.

Substances through which heat can not pass easily are **bad or poor conductors of heat**. Substances such as glass, wood, plastic, wax, *etc.*, are bad conductors of heat. Ebonite and asbestos are the **worst conductors of heat** and are called **insulators**.

The thermal conductivity of metals is very high because they contain a large number of free electrons which act as carriers of heat energy. Mercury is a good liquid conductor.

> Mercury is the only liquid which behaves like a metal.

Gases are the worst conductors of heat. Mica is a good conductor of heat but a bad conductor of electricity.

It has also been found that substances with low specific heat capacity (*e.g.* metals) are good conductors of heat energy. The higher the specific heat capacity of a substance, the lesser is the conductivity of the material. Such substances are generally bad conductors of heat.

RECAPITULATION

- The amount of heat absorbed by a substance depends on its mass, rise in its temperature and nature of the substance.
- One calorie is the amount of heat required to raise the temperature of one gram of water from 14·5°C to 15·5°C.
- The amount of heat required to raise the temperature of unit mass of a body through 1°C is called specific heat capacity of the substance of body.
- The amount of heat required to raise the temperature of entire mass of a substance through 1°C is called its heat capacity.
- The process of change of solid into liquid is called fusion.
- The process of change of vapour into a liquid is called condensation.
- The process of change of liquid into a solid is called freezing or solidification.
- The amount of heat required to convert a solid (a liquid) into liquid (a gas) without any rise in temperature is called its latent heat. If the mass is of unit amount, then the quantity is called specific latent heat.
- The process in which liquid changes into its vapour state, from its surface only, at all temperatures is called evaporation.
- The process in which a solid turns directly into its gaseous state is called sublimation.

TEST YOURSELF

A. Short Answer Questions :

1. State whether the following statements are *true* or *false*. Rewrite the correct statement if it is false.
   (a) Boiling takes place at all temperatures.
   (b) We feel refreshed under a tree during summers.
   (c) Evaporation takes place over the entire mass of the liquid.
   (d) Boiling water at 100°C produces more severe burns than steam at 100°C.
   (e) When heat energy is supplied to a substance, its intermolecular space increases.
   (f) According to the principle of calorimetry, heat lost by the hot object is equal to the heat gained by the cold object, when they are in contact.
   (g) In cold countries, juice bottles are kept in water to avoid their freezing.
   (h) The climate on hill station remains moderate due to land breeze and mountain breeze.
   (i) The unit of specific heat capacity is J kg^{-1} $°C^{-1}$.
   (j) The process of a gas converting into solid directly is called sublimation.

2. Fill in the blanks :
   (a) Water is used in a hot water bottle for..........
   (b) Water is used as .... in machines.
   (c) .... has highest specific heat capacity among liquids.
   (d) The cold air moving from sea to land is called ....
   (e) .... is a process just reverse of melting.
   (f) .... is process where a solid directly converts into gas.
   (g) The amount of heat required by a unit mass of a substance to change its state without any change in temperature is called ....
   (h) The SI unit of specific heat capacity is ....
   (i) The temperature at which a solid converts into a liquid is called its ....
   (j) Evaporation occurs only from the .... of the liquid.

3. Tick the appropriate answer :
   (a) The coolant used in a car engine is
   (i) alcohol (ii) benzene
   (iii) water (iv) petrol
   (b) The specific heat capacity of water is
   (i) 4250 J kg^{-1} $°C^{-1}$
   (ii) 4200 J kg^{-1} $°C^{-1}$
   (iii) 5000 J kg^{-1} $°C^{-1}$
   (iv) 4250 J kg^{-1} $°C^{-1}$
   (c) The amount of heat required to raise the temperature through 1°C is called
   (i) thermal energy (ii) calorie
   (iii) heat capacity
   (iv) specific heat capacity
   (d) During land breeze, there is a fall of pressure over the surface of
   (i) land (ii) water (sea)
   (iii) air (iv) none of these
   (e) The temperature at which a liquid gets converted into its vapour state is called its
   (i) melting point (ii) boiling point
   (iii) dew point (iv) freezing point
   (f) Rapid conversion of water into steam is an example of
   (i) evaporation (ii) freezing
   (iii) melting (iv) vapourisation
   (g) Evaporation takes place from
   (i) Surface of liquid
   (ii) Throughout the liquid
   (iii) From the mid-portion of the liquid
   (iv) None of these
   (h) Boiling takes place from
   (i) Surface of liquid

(ii) Throughout the liquid
(iii) From the mid-portion of liquid
(iv) None of these.

(i) The unit of specific latent heat is
(i) J kg^{-1}°C^{-1} (ii) J/kg°C Cal
(iii) J/kg (iv) None of these

(j) The unit of specific heat capacity is
(i) J kg^{-1}°C^{-1} (ii) J/kg°C Cal
(iii) J/kg (iv) None of these

4. Match the following :
(a) Specific latent heat of vapourisation of water — (i) 483 J kg^{-1} °C^{-1}
(b) Specific latent heat of fusion of ice — (ii) 100°C
(c) Boiling point of water — (iii) 0°C
(d) Melting point of ice — (iv) 2260000 J/kg
(e) Specific heat capacity of iron — (v) 336000 J/kg

5. Answer the following questions.
(a) Define one kilocalorie. How is it related to Joules?
(b) Express 336000 J in calories.
(c) Define specific heat capacity. What is its SI unit ?
(d) Define specific latent heat. What is its SI unit?
(e) Why is cooling produced during evaporation?
(f) Give three examples from daily life to show that evaporation produces cooling effect.

6. Solve the following :
(a) 1500 cal of heat is supplied to 200 g of water. Find the rise in temperature. Specific heat capacity of water is 1 cal g^{-1} °C^{-1}. **Ans.** 7·5°C
(b) 500 g of water at 100°C is mixed with 300 g of water at 20°C. Specific heat capacity of water is 4·2 J g^{-1}°C^{-1}. Find the final temperature of the mixture. **Ans.** 70°C
(c) Calculate the heat energy liberated from 40 g of steam so as to form water at 100°C (Sp. latent heat of vapourisation of steam is 3360 J/g). **Ans.** 134400 J
(d) Calculate the specific latent heat of fusion of a solid if 20 g of it requires 8400 J of heat energy for melting. **Ans.** 420 J g^{-1}
(e) A brass road of 0·2 kg at 100°C is dropped into 0·5 kg of water at 20°C. The final temperature of the mixture is 23°C. Calculate the specific heat capaicty of brass (Given specific heat capacity of water = 4200 J/kg°C). **Ans.** 409·09 J/kg°C
(f) 2 litres of water at 80°C is poured into a plastic bucket containing 10 litres of water at 20°C. What is the final temperature of water (Density of water 1 kg/litre) (neglect the heat gained by bucket and specific heat capacity of water = 1 cal/gm°C). **Ans.** 30°C

B. Long Answer Questions :

1. The specific heat capacity of water is 4200 J/kg°C. What does this statement mean ? Give specific heat capacities of kerosene and mercury.
2. How does the high specific heat capacity of water help in the formation of land-breeze and sea-breeze ?
3. What do you mean by the change of state ? Show the complete cycle of change of state.
4. Differentiate between melting point and boiling point, giving atleast one example in each case.
5. Describe the process of condensation and sublimation with examples.
6. What are the consequences of high specific latent heat capacity of steam ?
7. What are conductors and insulators ? Give examples. Why solids conduct heat better than liquids or gases ?