Chapter 11

Calorimetry

(E) Heat

SYLLABUS

(i) Calorimetry : Meaning, Specific heat capacities, Principle of method of mixture, Numerical problems on specific heat capacity using heat loss and gain and the method of mixtures.

Scope of syllabus : Heat and its units (calorie, joule), temperature and its units (°*C*, K); Thermal (heat) capacity $C'=Q/\Delta T$. (S.I. unit of C'), Specific heat capacity $c = Q/m\ \Delta T$; (S.I. unit of c), Mutual relations between heat capacity and specific heat capacity, Values of c for some common substances (ice, water and copper). Principle of method of mixtures including mathematical statement. Natural phenomena involving specific heat; consequences of high sp. heat of water. Simple numerical problems.

(ii) Latent heat; loss and gain of heat involving change of state for fusion only.

Scope of syllabus : Change of phase (state); heating curve for water; latent heat; sp latent heat of fusion (S.I. unit). Simple numerical problems. Common physical phenomena involving latent heat of fusion.

(A) HEAT CAPACITY, SPECIFIC HEAT CAPACITY AND ITS MEASUREMENT

11.1 CONCEPT OF HEAT

We know that each substance is made up of molecules. The molecules in a substance are in a state of random motion and each molecule exerts a force of attraction on the other molecules. Thus molecules possess kinetic energy due to their random motion and potential energy due to the molecular attractive forces. The sum of the potential energy and kinetic energy is called their *internal energy. The total internal energy of molecules of a substance is called its heat energy.*

A hot body has more internal energy than an identical cold body. When a hot body is kept in contact with a cold body, the cold body warms up, while the hot body cools down. *i.e.*, the internal energy of the cold body increases, while that of the hot body decreases. Thus there is a flow of internal energy from the hot body to the cold body when they are kept in contact. The energy which flows from the hot body to the cold body is called the *heat energy* or simply the *heat.* Thus

Heat is the internal energy of molecules constituting the body. It flows from a hot body to a cold body when they are kept in contact.

Like all other forms of energy, heat is also a measurable quantity. The measurement of the quantity of heat is called *calorimetry*.

Units of heat

Like other forms of energy, the S.I. unit of heat is *joule* (symbol J).

The other most commonly used unit of heat is *calorie* (symbol cal). It is defined as follows :

One calorie is the quantity of heat energy required to raise the temperature of 1 g *of water through* 1°C.

In the above definition, it has been assumed that the heat energy required to raise the temperature of 1 g of water through 1°C at each *initial* temperature is same. However this is not true due to non-uniform thermal expansion of water. Hence the precise definition of calorie (which is also called 15°C calorie) is given as follow :

One calorie is the heat energy required to raise the temperature of 1 g *of water from* 14·5°C *to* 15·5°C.

The unit calorie is related to the S.I. unit joule as follows :

1 calorie (or 1 cal) = 4·186 J or 4·2 J nearly* ...(11.1)

Sometimes, calorie is a smaller unit of heat, so we use a bigger unit called the kilo-calorie (symbol kcal), where

1 kilo-calorie = 1000 calorie = 4200 J nearly.

One kilo-calorie is the heat energy required to raise the temperature of 1 kg *of water from* 14·5°C *to* 15·5°C.

The unit kilo-calorie is generally used for measuring the energy value of foods.

11.2 CONCEPT OF TEMPERATURE

On keeping a hot body in contact with a cold body, heat flows from the hot body to the cold body due to which the hot body gets cooled, while the cold body gets warmed.

The body which imparts heat is said to be at a higher temperature than the body which receives heat. Thus, *temperature determines the direction of flow of heat.*

When a body receives heat energy, the particles constituting the body start vibrating more vigorously and so its temperature rises provided its physical state or dimensions remain unchanged.

Thus temperature is defined as below.

Temperature is a parameter which tells the thermal state of a body (i.e., the degree of hotness or coldness of body). It determines the direction of flow of heat when two bodies at different temperatures are placed in contact.

If there is no transfer of heat between the two bodies placed in contact, they are said to be at same temperature, but it does not mean that they have equal amount of heat in them. In fact, temperature alone does not tell us the quantity of heat energy contained in a body. Experimentally, we find that by imparting the same quantity of heat energy to different bodies, they get heated to different temperatures. *The amount of heat energy contained in a body depends on mass, temperature and the material of body.*

Unit of temperature

The S.I. unit of temperature is kelvin (symbol K). The other most common unit of temperature is degree celsius (symbol °C). They are related as :

$$\text{T K} = 273 + t\ °\text{C} \qquad ...(11.2)$$

or more precisely, T K = 273·15 + t °C

Thus by adding 273 (or 273·15) to the temperature in degree celsius, we get the temperature in kelvin. The zero of the kelvin scale (called absolute zero or 0 K) is the temperature at which the molecular motion ceases. It is equal to –273°C (or –273·15°C).

Thus *a degree (or temperature difference) is same on both the celsius and kelvin scales i.e.,*

$$\Delta t\ °\text{C} = \Delta \text{T K} \qquad ...(11.3)$$

11.3 FACTORS AFFECTING THE QUANTITY OF HEAT ABSORBED TO INCREASE THE TEMPERATURE OF A BODY

The quantity of heat energy absorbed to increase the temperature of a body depends on *three* factors : (1) mass of the body, (2) the increase in temperature of the body, and (3) the material (or substance) of the body.

Experimentally it is observed that

(1) Objects with different mass made from the same substance absorb different amounts of heat energy to raise their temperature by the same amount. For example, to raise the

* For calculations, we generally take 1 cal = 4·2 J

temperature of 1 kg of water by 1°C, heat energy absorbed is 1 kcal, while to raise the temperature of 2 kg of water by 1°C is 2 kcal. Thus the amount of heat energy absorbed is directly proportional to the mass of the object *i.e.*, $Q \propto m$. ...(i)

(2) Objects with equal mass made from the same substance absorb different amounts of heat energy to raise their temperature by different amounts. For example, to raise the temperature of 1 kg of water by 1°C, heat energy absorbed is 1 kcal, while to raise the temperature of same mass of water by 2°C is 2 kcal. Thus the amount of heat energy absorbed is directly proportional to the rise in temperature Δt *i.e.*, $Q \propto \Delta t$ ...(ii)

(3) Objects of same mass but made from different substances absorb different amounts of heat energy to raise their temperature by the same amount. For example, if equal mass of water and copper are heated through 1°C, the amount of heat absorbed by water is nearly ten times the amount of heat absorbed by copper. Thus, the amount of heat energy absorbed depends on the substance of the object which is expressed in terms of its *specific heat capacity c*.

From the above observations (i) and (ii),

$$Q \propto m \text{ and } Q \propto \Delta t$$

or $$Q = c\, m\, \Delta t \qquad ...(11.4)$$

where c is the constant of proportionality which is called the *specific heat capacity* of the substance. It is the characteristic of the substance and is different for different substances.

11.4 DIFFERENCE BETWEEN HEAT AND TEMPERATURE

Heat	Temperature
1. Heat is a form of internal energy obtained due to random motion and attractive force of molecules in a substance.	1. Temperature is a quantity which determines the direction of flow of heat on keeping the two bodies at different temperatures in contact.
2. The S.I. unit of heat is joule (J).	2. The S.I. unit of temperature is kelvin (K).
3. The amount of heat contained in a body depends on mass, temperature and substance of body.	3. The temperature of a body depends on the average kinetic energy of its molecules due to their random motion.
4. Heat is measured by the principle of calorimetry.	4. Temperature is measured by a thermometer.
5. Two bodies having same quantity of heat may differ in their temperature.	5. Two bodies at same temperature may differ in the quantities of heat contained in them.
6. When two bodies are placed in contact, the total amount of heat is equal to the sum of heat of individual body.	6. When two bodies at different temperatures are placed in contact, the resultant temperature is a temperature in between the two temperatures.

11.5 THERMAL (OR HEAT) CAPACITY ($C' = Q / \Delta T$)

From our everyday experience we find that different bodies require different amounts of heat energy for equal rise in their temperature. This property of a body is expressed in terms of its *thermal* (or *heat*) *capacity*. The heat capacity of a body is defined as follows :

The heat capacity of a body is the amount of heat energy required to raise its temperature by 1°C (or 1 K).

It is denoted by the symbol C'. Thus,

$$\text{Heat capacity } C' = \frac{\text{amount of heat energy supplied}}{\text{rise in temperature}} \qquad ...(11.5)$$

If on imparting an amount of heat Q to a body, its temperature rises through Δt °C (or Δt K), then

$$\text{Heat capacity of the body } C' = \frac{Q}{\Delta t} \qquad ...(11.6)$$

Units of heat capacity

The S.I. unit of heat capacity is *joule per kelvin* (or J K^{-1}). It is also written as joule per degree C (or J $°C^{-1}$)*.

The other common units of heat capacity are cal $°C^{-1}$ (or cal K^{-1}) and kcal $°C^{-1}$ (or kcal K^{-1}). They are related as

$$1 \text{ kcal } °C^{-1} = 1000 \text{ cal } °C^{-1}$$
and $$1 \text{ cal } K^{-1} = 4{\cdot}2 \text{ J } K^{-1}$$

Note : If the heat capacity of a vessel is 30 J K^{-1}, it means that 30 J heat energy is required to raise the temperature of that vessel by 1 K (or 1°C).

11.6 SPECIFIC HEAT CAPACITY

Heat capacity of a body when expressed for unit mass is called its *specific heat capacity*. It is denoted by the symbol *c*. Thus *specific heat capacity of a substance is defined as the heat capacity per unit mass of a body of that substance, i.e.,*

$$\text{Specific heat capacity } c = \frac{\text{Heat capacity of body } C'}{\text{Mass of the body } m} \quad \text{... (11.7)}$$

From eqn. (11.2), $C' = \frac{Q}{\Delta t}$

$$\therefore \text{ Specific heat capacity } c = \frac{Q}{m \times \Delta t} \quad \text{...(11.8)}$$

In other words, we can define specific heat capacity as follows :

The specific heat capacity of a substance is the amount of heat energy required to raise the temperature of unit mass of that substance through 1°C (or 1 K). i.e.,

Specific heat capacity

$$c = \frac{\text{amount of heat energy supplied}}{\text{mass} \times \text{rise in temperature}} \quad \text{...(11.9)}$$

* Since S.I. unit of temperature is kelvin (K), so it is more appropriate to write the S.I. unit of heat capacity as J K^{-1} instead of J $°C^{-1}$.

From eqn. (11.9), heat energy required to raise the temperature of a body of mass *m* kg by Δt K is given as :

$$Q = m \times c \times \Delta t \text{ joule} \quad \text{..(11.10)}$$

Units of specific heat capacity

The S.I. unit of specific heat capacity is *joule per kilogram per kelvin* (or J kg^{-1} K^{-1}) or joule per kilogram per degree celsius (or J kg^{-1} $°C^{-1}$).

The other units of specific heat capacity are cal g^{-1} $°C^{-1}$ and kcal kg^{-1} $°C^{-1}$. These units are related as :

$$1 \text{ cal } g^{-1} \, °C^{-1} = 1 \text{ kcal } kg^{-1} \, K^{-1}$$
$$= 4{\cdot}2 \times 10^3 \text{ J } kg^{-1} \, K^{-1}$$

Note : If specific heat capacity of copper is 0·4 J g^{-1} K^{-1}, it means that the heat energy required to raise the temperature of 1 g of copper by 1 K (or 1°C) is 0·4 J.

Relationship between the heat capacity and specific heat capacity

From eqn. (11.7).

$$\text{Heat capacity } C' = \text{mass } m \times \text{specific heat capacity } c \quad \text{..(11.11)}$$

The above eqn. (11.11) relates the heat capacity C' of a body to the specific heat capacity *c* of its substance.

11.7 DISTINCTION BETWEEN THE HEAT CAPACITY AND SPECIFIC HEAT CAPACITY

Heat capacity	Specific heat capacity
1. It is the amount of heat energy required to raise the temperature of entire body by 1°C.	1. It is the amount of heat energy required to raise the temperature of unit mass of the body by 1°C.
2. It depends both on the substance and mass of the body. More the mass of the body, more is its heat capacity.	2. It does not depend on the mass of the body, but it is the characteristic property of the substance of the body.
3. Heat capacity $C' = \frac{Q}{\Delta t}$ = mass m × specific heat capacity c	3. Specific heat capacity $c = \frac{Q}{m\,\Delta t} = \frac{\text{heat capacity } C'}{\text{mass } m}$
4. Its unit is J K^{-1}.	4. Its unit is J kg^{-1} K^{-1}.

11.8 SPECIFIC HEAT CAPACITY OF SOME COMMON SUBSTANCES

The specific heat capacity of a substance is its characteristic property. It is different for different substances. Usually a good conductor has a low specific heat capacity, while a bad conductor has a high specific heat capacity. If we heat *equal mass* of two different substances on the same burner so that the rate of heat supply is same, we notice that after the same time interval, the rise in temperature for the two substances is different. This is because of their different specific heat capacities. *The substance with low specific heat capacity shows a rapid and high rise in temperature thus it is a better conductor of heat than the substance with high specific heat capacity which shows a slow and small rise in temperature.*

Water has an unusually high specific heat capacity (= 4200 J kg^{-1} K^{-1}).*

> *Note :* The specific heat capacity of the same substance is different in its different phases. The specific heat capacity of water is 4200 J kg^{-1} K^{-1}, of ice is 2100 J kg^{-1} K^{-1} and of steam is 460 J kg^{-1} K^{-1}.

The specific heat capacity is maximum, equal to 14630 J kg^{-1} K^{-1} for hydrogen.

Specific heat capacity of some common substances

Substance	Specific heat capacity	
	in J kg^{-1} K^{-1}	in cal g^{-1} °C^{-1}
1. Lead	130	0·031
2. Mercury	139	0·033
3. Brass	380	0·092
4. Zinc	391	0·093
5. Copper	399	0·095
6. Iron	483	0·115
7. Glass (flint)	504	0·12
8. Aluminium	882	0·21
9. Kerosene oil	2100	0·50
10. Ice	2100	0·50
11. Sea water	3900	0·95
12. Water	4180	1·0

* More precisely the specific heat capacity of water is 4180 J kg^{-1} K^{-1}. But for convenience, its value is taken as 4200 J kg^{-1} K^{-1}.

11.9 CALORIMETER

A calorimeter is a cylindrical vessel which is used to measure the amount of heat gained (or lost) by a body when it is mixed with the other body. It is shown in Fig. 11.1. It is made up of a thin sheet of copper because (i) copper is a good conductor of heat, so the vessel soon acquires the temperature of its contents, and (ii) copper has the low specific heat capacity so the heat capacity of calorimeter is low and the amount of heat energy taken by the calorimeter from its contents to acquire the temperature of its contents, is very small. The outer and inner surfaces of vessel are polished so as to reduce the loss of heat due to *radiation*. For insulation, it is placed inside a wooden jacket. The space between the calorimeter and the jacket is filled with some poor conductor such as wool, cotton, etc. to avoid the heat loss by *conduction*. It is covered with a wooden lid to avoid the heat loss by *convection*. The lid has two holes, one for the stirrer (used to mix its contents properly) and the other for the thermometer (to measure the temperature of its contents).

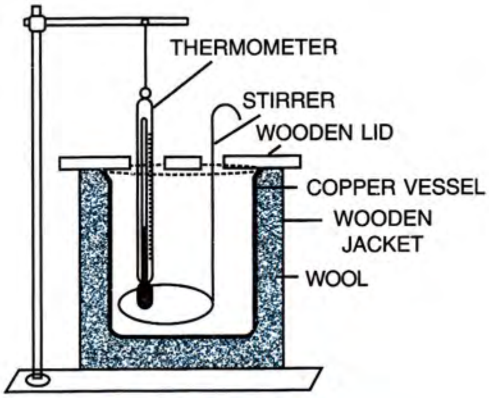

Fig. 11.1 Calorimeter

11.9 PRINCIPLE OF METHOD OF MIXTURES (OR PRINCIPLE OF CALORIMETRY)

When a hot body is mixed (or is kept in contact) with a cold body, heat energy passes from the hot body to the cold body, till both the bodies attain the same temperature. If no heat energy is lost to the surroundings (*i.e.*, if the system is perfectly insulated), then

Heat energy lost by the hot body = Heat energy gained by the cold body ..(11.12)

This is called the *principle of method of mixtures* (or the *principle of calorimetry*). This principle is based on the law of conservation of energy.

Mathematical statement : Let a substance A of mass m_1, specific heat capacity c_1 at a higher temperature t_1 *be mixed with* another substance B of mass m_2, specific heat capacity c_2 at a lower temperature t_2 (*i.e.*, $t_2 < t_1$). If the final temperature of mixture becomes t, then

Fall in temperature of the substance A = $t_1 - t$

Rise in temperature of substance B = $t - t_2$

Heat energy lost by A = $m_1 \times c_1 \times$ fall in temperature

$= m_1 c_1 (t_1 - t)$ .... (i)

Heat energy gained by B = $m_2 \times c_2 \times$ rise in temperature

$= m_2 c_2 (t - t_2)$ .... (ii)

If no heat energy is lost to the surroundings, then by the principle of method of mixtures,

Heat energy lost by A = Heat energy gained by B

or $$m_1 c_1 (t_1 - t) = m_2 c_2 (t - t_2) \quad ...(11.13)$$

The above expression (11.13) can be used to calculate the unknown quantity.

Note : If calorimeter (or a vessel) is used for mixing the two substances A and B, we must take into account the heat energy gained by the calorimeter also. If mass of calorimeter is M and specific heat capacity of its substance is c, it wil gain heat energy = Mc ($t_1 - t$). Then eqn. (11.3) will take the form

$$m_1c_1(t_1 - t) = m_2c_2(t - t_2) + Mc\ (t - t_2)$$
$$= (m_2c_2 + Mc)\ (t - t_2) \quad ..(11.14)$$

11.10 NATURAL PHENOMENA AND CONSEQUENCES OF HIGH SPECIFIC HEAT CAPACITY OF WATER

Some consequences of high specific heat capacity of water (= 4200 J kg^{-1} K^{-1})are given below.

(1) **The climate near the sea shore is moderate:** The specific heat capacity of water is very high (= 1000 cal kg^{-1} °C^{-1} or 4200 J kg^{-1} K^{-1}). *It is about five times as high as that of sand.* Hence the heat energy required for the same rise in temperature by a certain mass of water will be nearly five times than that required by the same mass of sand. Similarly, a certain mass of water will impart nearly five times more heat energy than that given by the same mass of sand for the same fall in temperature. As such, *sand* (or *earth*) *gets heated* or *cooled more rapidly as compared to water under similar conditions.* Thus, near the sea shore, there becomes a large difference in temperature between the land and sea due to which convection air currents are set up. The cold air blows from the land towards the sea during night (*i.e.*, land breeze) and during the day, cold air blows from the sea towards the land (*i.e.*, sea breeze). *These breezes make the climate near the sea shore moderate.*

(2) **Hot water bottles are used for fomentation :** The reason is that water does not cool quickly due to its high specific heat capacity, so a hot water bottle provides more heat energy for fomentation over a longer period.

(3) **Water is used as an effective coolant :** By allowing water to flow in pipes around the heated parts of a machine, heat energy from such parts is removed. Water in pipes can extract more heat from the surroundings without much rise in its temperature because of its high specific heat capacity. This is why radiators in car and generator are filled with water.

(4) **In cold countries, water is used as heat reservoir for wine and juice bottles to avoid their freezing :** The reason is that water due to its high specific heat capacity can impart a large amount of heat before reaching to its freezing point. Hence bottles kept in water remain warm and they do not

freeze even when the surrounding temperature falls considerably.

(5) Farmers fill their fields with water to protect the crops from frost : On a cold winter night, if the atmospheric temperature falls below 0°C, water in the fine capillaries of plants will freeze, so the veins will burst due to the increase in volume of water on freezing. As a result, plants will die and the crop will get destroyed. In order to save crop on such cold nights, farmers fill their fields with water because water has a high specific heat capacity, so it does not allow the temperature in the surrounding area of plants to fall up to 0°C.

(6) All plants and animals have a high content of water in their bodies : All plants and animals have nearly 80 to 90% of water in their bodies. This water, because of its high specific heat capacity, maintains nearly same temperature of their body in all seasons.

11.11 SOME EXAMPLES OF HIGH AND LOW HEAT CAPACITY

(1) The base of a cooking pan is made thick : By making the base of the cooking pan thick, its heat capacity becomes large due to which it gets heated slowly and it imparts sufficient heat energy at a slow rate to the food for its proper cooking and after cooking, it keeps the food warm for a long time.

(2) The base of an electric iron is made thick and heavy : By doing so, the heat capacity of base of the electric iron becomes large and it remains hot for a long duration even after the current is switched off.

(3) The vessel used for measurement of heat (*i.e.*, calorimeter) is made of a thin sheet of copper : The reason is that the specific heat capacity of copper is low and by making the vessel thin, its mass and hence its heat capacity becomes low so that it takes a negligible amount of heat from its contents to attain the temperature of the contents.

EXAMPLES

1. An iron ball requires 5000 J heat energy to raise its temperature by 10°C. Calculate the heat capacity of the iron ball.

Given, $Q = 5000$ J, $\Delta t = 10°C = 10$ K

$$\text{Heat capacity } C' = \frac{\text{heat energy required } Q}{\text{rise in temperature } \Delta t}$$

$$\text{or} \quad C' = \frac{5000 \text{ J}}{10 \text{ K}} = \mathbf{500 \text{ J K}^{-1}}$$

2. Calculate the heat energy required to raise the temperature of 2 kg of water from 10°C to 50°C. Specific heat capacity of water is 4200 J kg^{-1} K^{-1}.

Given, $m = 2$ kg, $c = 4200$ J kg^{-1} K^{-1},

$\Delta t = (50 - 10)°C = 40°C = 40$ K

$$\text{Heat energy required } Q = m\,c\,\Delta t$$
$$= 2 \times 4200 \times 40$$
$$= \mathbf{336000 \text{ J}}$$

3. A metal piece of mass 50 g at 27°C requires 2400 J of heat energy in order to raise its temperature to 327°C. Calculate the specific heat capacity of the metal.

Given : $m = 50$ g $= 0{\cdot}05$ kg, heat energy required $Q = 2400$ J, rise in temperature $\Delta t = (327 - 27)°C = 300°C = 300$ K

Specific heat capacity

$$c = \frac{\text{heat energy } Q}{\text{mass } m \times \text{rise in temperature } \Delta t}$$
$$= \frac{2400 \text{ J}}{0{\cdot}05 \text{ kg} \times 300 \text{ K}} = \mathbf{160 \text{ J kg}^{-1} \text{ K}^{-1}}$$

4. Some heat energy is given to 120 g of water and its temperature rises by 10 K. When the same amount of heat energy is given to 60 g of oil, its temperature rises by 40 K. The specific heat capacity of water is 4200 J kg^{-1} K^{-1}. Calculate : (i) the amount of heat energy in joule given to water, and (ii) the specific heat capacity of oil.

(i) Heat energy given to water

= mass of water × specific heat capacity of water × rise in temperature

$$= \frac{120}{1000} \text{kg} \times 4200 \text{ J kg}^{-1} \text{ K}^{-1} \times 10 \text{ K} = \mathbf{5040 \text{ J}}$$

(ii) Heat energy given to oil = 5040 J (Given).

Let c J kg^{-1} K^{-1} be the specific heat capacity of oil, then

$$c = \frac{\text{amount of heat energy given to oil}}{\text{mass of oil} \times \text{rise in temperature}}$$

$$= \frac{5040\text{J}}{\left(\frac{60}{1000}\text{kg}\right) \times 40\text{K}} = \mathbf{2100\ J\ kg^{-1}\ K^{-1}}$$

5. An electric heater of power 1000 W raises the temperature of 5 kg of a liquid from 25°C to 31°C in 2 minutes. Calculate : (i) the heat capacity, and (ii) the specific heat capacity of liquid.

Time t = 2 minutes = 2 × 60 s = 120 s

Rise in temperature ΔT = (31 – 25)°C = 6°C = 6 K

Mass of liquid m = 5 kg

Energy supplied by the heater = power × time

$$Q = 1000\text{ W} \times 120\text{ s} = 1{\cdot}2 \times 10^5\text{ J}$$

(i) Heat capacity $C' = \frac{\text{energy supplied } Q}{\text{rise in temperature } \Delta T}$

$$= \frac{1{\cdot}2 \times 10^5\text{ J}}{6\text{ K}} = \mathbf{2 \times 10^4\ J\ K^{-1}}$$

(ii) Specific heat capacity

$$c = \frac{\text{heat capacity } (C')}{\text{mass } (m)} = \frac{2 \times 10^4\text{ J K}^{-1}}{5\text{ kg}}$$

$$= \mathbf{4 \times 10^3\ J\ kg^{-1}\ K^{-1}}$$

6. A bucket contains 8 kg of water at 25°C. 2 kg of water at 80°C is poured into it. Neglecting the heat energy absorbed by the bucket, calculate the final temperature of water.

Let final temperature of water be t °C.

Fall in temperature of hot water = (80 – t)°C

Heat energy lost by hot water = 2 × c × (80 – t) ...(i)

Rise in temperature of cold water = (t – 25)°C

Heat energy gained by cold water = 8 × c × (t – 25) ...(ii)

where c is the specific heat capacity of water.

Neglecting the heat energy absorbed by the bucket, heat energy lost by hot water

= heat energy gained by cold water

$$2 \times c \times (80 - t) = 8 \times c \times (t - 25)$$

or $2(80 - t) = 8(t - 25)$

or $160 - 2t = 8t - 200$

or $10t = 360$

or $t = \frac{360}{10} = \mathbf{36°C}$

Thus, final temperature of water will be **36°C**.

7. 40 g of water at 60°C is poured into a vessel containting 50 g of water at 20°C. The final temperature of mixture is 30°C. Taking the specific heat capacity of water as 4·2 J g^{-1} K^{-1}, calculate the heat capacity of the vessel.

Let heat capacity of vessel be C' J K^{-1}

Heat energy given by hot water

= mass of hot water × specific heat capacity × fall in temperature

= 40 × 4·2 × (60 – 30) = 5040 J ...(i)

Heat energy taken by cold water

= mass of cold water × specific heat capacity × rise in temperature

= 50 × 4·2 × (30 – 20) = 2100 J ...(ii)

Heat energy taken by vessel

= heat capacity of vessel × rise in temperature

= C' × (30 – 20) = 10 C' J ...(iii)

If there is no loss of heat energy,

Heat energy given by hot water

= Heat energy taken by cold water + Heat energy taken by vessel

or $5040 = 2100 + 10C'$

or $10C' = 2940$

or $C' = 294$ J K^{-1}.

Thus, heat capacity of vessel = **294 J K^{-1}**.

8. A metal piece of mass 20 g is heated to a constant temperature of 100°C. Then it is added in a calorimeter of mass 50 g and specific heat capacity 0·42 J g^{-1} K^{-1}, containing 50 g of water at 20°C. After stirring the water, the highest temperature recorded is 22°C. Calculate the specific heat capacity of metal.
Specific heat capacity of water = 4·2 J g^{-1} K^{-1}.

Let specific heat capacity of metal be c J g^{-1} K^{-1}.

Heat energy given by metal

= 20 × c × (100 – 22) = 1560 c J ..(i)

Heat energy taken by calorimeter

= 50 × 0·42 × (22 – 20)J = 42 J

Heat energy taken by water

= 50 × 4·2 × (22 – 20) J = 420 J

Total heat energy taken by calorimeter and water

= 42 + 420 = 462 J ..(ii)

If there is no loss of heat energy,

heat energy given by metal = total heat energy taken by calorimeter and water

or $1560\,c = 462$

or $c = \frac{462}{1560} = $ **0·3 J g^{-1} K^{-1}**

9. A hot iron ball of mass 0·2 kg is added into 0·5 kg of water at 10°C. The resulting temperature is 30°C. Calculate the temperature of hot ball. Specific heat capacity of iron = 336 J kg^{-1} K^{-1} and specific heat capacity of water = 4·2 × 10^3 J kg^{-1} K^{-1}.

Let temperature of hot ball be t °C

Fall in temperature of ball = $(t - 30)$°C

Rise in temperature of water = (30 – 10)°C

Heat energy given by ball = $0·2 \times 336 \times (t - 30)$ J ...(i)

Heat energy taken by water

$= 0·5 \times (4·2 \times 10^3) \times (30 - 10)$ J ...(ii)

Assuming that there is no loss of heat energy,

heat energy given by ball

= heat energy taken by water

$0·2 \times 336 \times (t - 30) = 0·5 \times 4·2 \times 10^3 \times 20$

or $67·2\,t - 2016 = 42000$

or $67·2\,t = 42000 + 2016 = 44016$

or $t = \frac{44016}{67·2} = $ **655 °C**

Thus, temperature of hot ball was **655 °C.**

10. What mass of a liquid A of specific heat capacity 0·84 J g^{-1} K^{-1} at a temperature 40°C must be mixed with 100 g of a liquid B of specific heat capacity 2·1 J g^{-1} K^{-1} at 20°C, so that the final temperature of mixture becomes 32°C ?

Let m g of liquid A be required.

Fall in temperature of liquid A = (40 – 32) °C

Rise in temperature of liquid B = (32 – 20) °C

Heat energy given by m g of liquid A

$= m \times 0·84 \times (40 - 32)$ J ..(i)

Heat energy taken by 100 g of liquid B

$= 100 \times 2·1 \times (32 - 20)$ J ..(ii)

Assuming that there is no heat loss,

∴ heat energy given by A = heat energy taken by B

or $m \times 0·84 \times 8 = 100 \times 2·1 \times 12$

∴ $m = \frac{100 \times 2·1 \times 12}{0·84 \times 8} = $ **375 g**

11. Water initially at 20°C at a height of 1·68 km above the ground, falls down on ground. Taking the specific heat capacity of water to be 4200 J kg^{-1} K^{-1}, find the final temperature of water on reaching the ground. Take g = 10 m s^{-2}.

Given, initial temperature = 20°C, g = 10 m s^{-2},

$h = 1·68$ km $= 1·68 \times 10^3$ m and $c = 4200$ J kg^{-1} K^{-1}.

Initially, water at a height h has the potential energy mgh stored in it which changes into the kinetic energy during the fall. On reaching the ground, water has the kinetic energy equal to the initial potential energy (equal to mgh) which on striking the ground changes into the heat energy $mc\Delta t$, if Δt is the rise in temperature of water. Assuming that whole of the kinetic energy of water changes into the heat energy and there is no loss of energy,

$$mgh = mc\Delta t$$

or rise in temperature $\Delta t = \frac{gh}{c}$

$= \frac{10 \times (1·68 \times 10^3)}{4200} = 4$°C

Hence final temperature of water on reaching the ground = initial temperature + rise in temperature

= 20°C + 4°C = **24°C.**

EXERCISE–11(A)

1. Define the term heat.
2. Name the S.I. unit of heat. **Ans.** joule (J)
3. Define the term calorie. How is it related to joule ?
4. Define one kilo-calorie of heat.
5. Define temperature and name its S.I. unit.
6. Differentiate between heat and temperature.
7. Define calorimetry.
8. Define the term heat capacity and state its S.I. unit.
9. Define the term specific heat capacity and state its S.I. unit.
10. How is heat capacity of a body related to specific heat capacity of its substance ?
11. State *three* differences between the heat capacity and specific heat capacity.

12. Name a liquid which has the highest specific heat capacity. **Ans.** Water

13. Write the approximate value of specific heat capacity of water in S.I. unit. **Ans.** 4200 J kg^{-1} K^{-1}

14. What do you mean by the following statements :
(i) the heat capacity of a body is 50 J K^{-1} ?
(ii) the specific heat capacity of copper is 0·4 J g^{-1} K^{-1} ?

15. Specific heat capacity of a substance A is 3·8 J g^{-1} K^{-1} and of substance B is 0·4 J g^{-1} K^{-1}. Which substance is a good conductor of heat ? How did you arrive at your conclusion ?
Ans. B. for the same heat energy and same mass, the rise in temperature of B will be more hence, B is a good conductor of heat.

16. Name *three* factors on which the heat energy absorbed by a body depends and state how does it depend on them.

17. Write the expression for the heat energy Q received by m kg of a substance of specific heat capacity c J kg^{-1} K^{-1} when it is heated through Δt °C. **Ans.** $Q = mc\Delta t$ J

18. Same amount of heat is supplied to two liquids A and B. The liquid A shows a greater rise in temperature. What can you say about the heat capacity of A as compared to that of B ?
Ans. heat capacity of A is less than that of B

19. Two blocks P and Q of different metals having their mass in the ratio 2 : 1 are given same amount of heat. Their temperature rises by same amount. Compare their specific heat capacities. **Ans.** 1 : 2

20. What is the principle of method of mixture ? What other name is given to it ? Name the law on which this principle is based.

21. A mass m_1 of a substance of specific heat capacity c_1 at temperature t_1 is mixed with a mass m_2 of other substance of specific heat capacity c_2 at a lower temperature t_2. Deduce the expression for the temperature of the mixture. State the assumption made, if any.
Ans. $\dfrac{m_1c_1t_1+m_2c_2t_2}{m_1c_1+m_2c_2}$
Assumption : There is no loss of heat energy.

22. Why do the farmers fill their fields with water on a cold winter night ?

23. Discuss the role of high specific heat capacity of water with reference to climate in coastal areas.

24. Water is used in hot water bottles for fomentation. Give a reason.

25. What property of water makes it an effective coolant ?

26. Give *one* example each where high specific heat capacity of water is used (i) as coolant, (ii) as heat reservoir.

27. A liquid X has specific heat capacity higher than the liquid Y. Which liquid is useful as (i) coolant in car radiators, and (ii) heat reservior to keep juice bottles without freezing ? **Ans.** (i) X (ii) X

28. What is a calorimeter ? Name the material of which it is made of. Give *two* reasons for using the material stated by you.

29. Why is the base of a cooking pan made thick and heavy ?

MULTIPLE CHOICE TYPE

1. The S.I. unit of heat capacity is :
(a) J kg^{-1} (b) J K^{-1}
(c) J kg^{-1} K^{-1} (d) cal °C^{-1} **Ans.** (b) J K^{-1}

2. The S.I. unit of specific heat capacity is :
(a) J kg^{-1} (b) J K^{-1}
(c) J kg^{-1} K^{-1} (d) kcal kg^{-1} °C^{-1}
Ans. (c) J kg^{-1} K^{-1}

3. The specific heat capacity of water is :
(a) 4200 J kg^{-1} K^{-1} (b) 420 J g^{-1} K^{-1}
(c) 0·42 J g^{-1} K^{-1} (d) 4·2 J kg^{-1} K^{-1}
Ans. (a) 4200 J kg^{-1} K^{-1}

NUMERICALS

1. By imparting heat to a body, its temperature rises by 15°C. What is the corresponding rise in temperature on kelvin scale ? **Ans.** 15 K

2. (a) Calculate the heat capacity of a copper vessel of mass 150 g if the specific heat capacity of copper is 410 J kg^{-1} K^{-1}.
(b) How much heat energy will be required to increase the temperature of the vessel in part (a) from 25°C to 35°C ?
Ans. (a) 61·5 J K^{-1}, (b) 615 J

3. A piece of iron of mass 2·0 kg has a heat capacity of 966 J K^{-1}. Find : (i) heat energy needed to warm it by 15°C, and (ii) its specific heat capacity in S.I. unit.
Ans. (i) 14490 J, (ii) 483 J kg^{-1} K^{-1}.

4. Calculate the amount of heat energy required to raise the temperature of 100 g of copper from 20°C to 70°C. Specific heat capacity of copper = 390 J kg^{-1} K^{-1}. **Ans.** 1950 J

5. 1300 J of heat energy is supplied to raise the temperature of 0·5 kg of lead from 20°C to 40°C. Calculate the specific heat capacity of lead. **Ans.** 130 J kg^{-1} K^{-1}

6. Find the time taken by a 500 W heater to raise the temperature of 50 kg of material of specific heat capacity 960 J kg^{-1} K^{-1} from 18°C to 38°C. Assume that all the heat energy supplied by heater is given to the material. **Ans.** 32 min.

7. An electric heater of power 600 W raises the temperature of 4·0 kg of a liquid from 10·0 °C to 15·0 °C in 100 s. Calculate : (i) the heat capacity of 4·0 kg of liquid, and (ii) the specific heat capacity of liquid.
Ans. (i) $1{\cdot}2 \times 10^4$ J K^{-1} (ii) 3×10^3 J kg^{-1} K^{-1}

8. 0·5 kg of lemon squash at 30°C is placed in a refrigerator which can remove heat at an average rate of 30 J s^{-1}. How long will it take to cool the lemon squash to 5°C ? Specific heat capacity of squash = 4200 J kg^{-1} K^{-1}.
Ans. 29 min 10 s

9. A mass of 200 g of a certain metal at 83°C is immersed in 300 g of water at 30°C. The final temperature is 33°C. Calculate the specific heat capacity of the metal. Assume that the specific heat capacity of water is 4·2 J g^{-1} K^{-1}.
Ans. 0·378 J g^{-1} K^{-1}

10. 45 g of water at 50°C in a beaker is cooled when 50 g of copper at 18°C is added to it. The contents are stirred till a final constant temperature is reached. Calculate the final temperature. The specific heat capacity of copper is 0·39 J g^{-1} K^{-1} and that of water is 4·2 J g^{-1} K^{-1}. State the assumption used.
Ans. 47°C
Assumption : There is no loss of heat.

11. 200 g of hot water at 80°C is added to 300 g of cold water at 10°C. Neglecting the heat taken by the container, calculate the final temperature of the mixture of water. Specific heat capacity of water = 4200 J kg^{-1} K^{-1}. **Ans.** 38°C

12. The temperature of 600 g of cold water rises by 15°C when 300 g of hot water at 50°C is added to it. What was the initial temperature of the cold water ? **Ans.** 5°C

13. 1·0 kg of water is contained in a 1·25 kW kettle. Calculate the time taken for the temperature of water to rise from 25°C to its boiling point 100°C. Specific heat capacity of water = 4·2 J g^{-1} K^{-1}.
Ans. 4 min 12 s

(B) CHANGE OF PHASE (STATE) AND LATENT HEAT

11.12 CHANGE OF PHASE (STATE)

There are *three* states (or phases) of matter namely *solid, liquid* and *gas*. The same matter can exist in all the three phases under different conditions of temperature and pressure. For example, ice (solid) when heated becomes water (liquid), which on further heating changes to steam (gas). Thus at one atmospheric pressure, water is found in all the three phases at different temperatures.

The process of change from one state to another *at a constant temperature* is called the change of phase. *It is brought about by the exchange of heat.* Fig. 11.2 represents the different processes of change of phase.

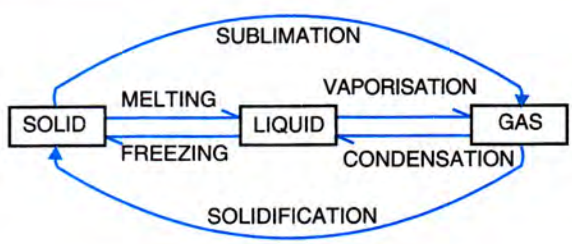

Fig. 11.2 Changes of phase

The change from solid to liquid phase is known as *melting,* while the reverse change from liquid to solid is called *freezing*.

The change from liquid to vapour is known as *vaporisation*, while the reverse change from gas (or vapour) to liquid is called *condensation* (or *liquefaction*).

The direct change from solid to vapour is called *sublimation* and the reverse change from vapour to solid is called *solidification*.

11.13 MELTING AND FUSION

The change from solid to liquid phase by the absorption of heat at a constant temperature is called melting.

*The constant temperature at which a solid changes to liquid is called the **melting point** of the solid.*

The reverse change from liquid to solid phase with the rejection of heat at a constant temperature is called *freezing* (or fusion) and the temperature at which a liquid freezes to solid is called its *freezing point*.

Thus heat energy is absorbed during melting and it is rejected during freezing at a constant temperature.

Example :

Melting
Absorption of heat
0°C ice ⇌ 0°C water
Rejection of heat
Freezing

Note : (1) For a pure substance, melting point and freezing point are identical.

(2) For a given mass of substance, the amount of heat energy absorbed during melting is same as that rejected during freezing.

11.14 HEATING CURVE OF ICE DURING MELTING

To draw the heating curve of ice, we perform the experiment as below.

Experiment : (1) Take a boiling test tube and fill it half with the ice chips. Insert a thermometer gently into the ice taking care that its bulb does not touch the walls of the test tube. Note the temperature of ice. It will be 0°C.

(2) Heat the bottom of the test tube slowly over a flame (or by immersing the test tube in a beaker containing hot water) and note the temperature after every half minute till the temperature of water formed after melting of ice, increases to 30°C.

Plot a graph for temperature against time by taking temperature on Y-axis and time on X-axis. The graph so obtained is shown in Fig. 11.3. It is called the *heating curve for ice*.

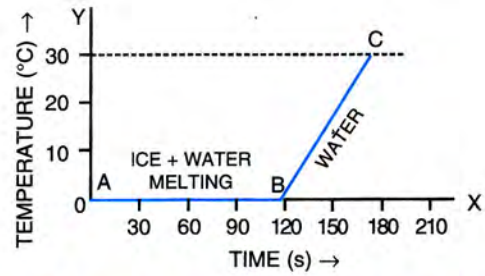

Fig. 11.3 Heating curve for ice initially at 0°C

From the graph, it is clear that the temperature of ice remains constant equal to 0°C in the part AB till the whole ice melts. The constant temperature at which the ice melts (*i.e.*, 0°C) is the melting point of ice. The heat supplied during this time is being used in melting the ice. After this, the temperature of water formed by the melted ice begins to rise from 0°C in the part BC.

11.15 CHANGE IN VOLUME ON MELTING

Most of the substances like lead, wax, etc. expand on melting, but some substances like ice contract on melting. For example, 1 g solid wax of volume 1·161 cm^3 on melting at 64°C becomes 1 g molten wax of volume 1·166 cm^3. On the other hand, 1 g ice of volume 1·091 cm^3 at 0°C on melting becomes 1 g water of volume 1 cm^3 at 0°C.

11.16 EFFECT OF PRESSURE ON THE MELTING POINT

The melting point of the substances which contract on melting (like ice) decreases by the increase in pressure. For example, the melting point of ice decreases by 0·0072°C for every one atmosphere rise in pressure.

On the other hand, *the melting point of the substances (such as wax, lead, etc.) which expand on melting, increases by the increase in pressure*.

11.17 EFFECT OF IMPURITIES ON THE MELTING POINT

The melting point of a substance decreases by the presence of impurities in it. The melting point of ice decreases from 0°C to – 22°C on mixing salt to it in proper proportion. This fact is utilised in making the *freezing mixture* by adding salt to ice. The freezing mixture is used in preparing '*kulphies*'.

11.18 VAPORISATION OR BOILING

The change from liquid to gas (or vapour) phase on absorption of heat at a constant temperature, is called vaporisation.

The particular temperature at which vaporisation occurs is called the ***boiling point*** *of liquid.*

Similarly, the change from vapour to liquid phase on rejection of heat at a constant temperature is called condensation (or liquefaction) and the particular temperature at which the condensation occurs is called the condensation point of vapour.

Note : For a pure substance, the boiling point and condensation point are identical.

Heat energy is absorbed at a constant temperature during vaporisation, while the same amount of heat energy is rejected during condensation at that temperature for the same mass of substance.

Example :

$$\text{100°C water} \underset{\text{Rejection of heat, Condensation}}{\overset{\text{Boiling, Absorption of heat}}{\rightleftharpoons}} \text{100°C steam}$$

11.19 HEATING CURVE FOR WATER

To obtain the heating curve of water, we perform the experiment as below.

Experiment : (1) Take some water at room temperature (say, 20°C) in a flask and suspend a thermometer in it.

(2) Heat the flask by keeping it over a bunsen burner and note its temperature after every half minute till the water starts boiling and the flask contains vapours. At this stage, bubbles are formed through out the water. This indicates that boiling occurs through out the volume of water.

Plot a graph for temperature against time, by taking the temperature on Y-axis and time on X-axis. The graph is shown in Fig. 11.4. It is called the *heating curve for water.*

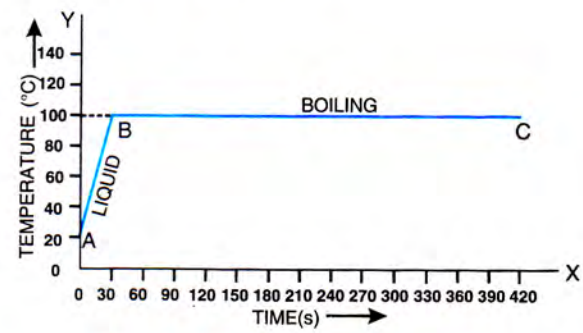

Fig 11.4 Heating curve for water

From the graph, it is observed that initially at A, water is at 20°C (room temperature) and then with the absorption of heat energy, the temperature of water rises continuously in the part AB where it is in the liquid phase. At B, the boiling starts and the temperature does not rise further in part BC, *although heat energy is being continuously absorbed.* The part BC is thus parallel to the time axis and represents the boiling of water. The particular temperature (= 100°C) at the point B is the boiling point of water.

11.20 CHANGE IN VOLUME ON BOILING

All liquids expand on boiling. For example, 1 cm^3 of water at 100°C becomes 1760 cm^3 of steam at 100°C.

11.21 EFFECT OF PRESSURE ON THE BOILING POINT

The boiling point of a liquid increases with the increase in pressure and decreases with the decrease in pressure.

The boiling point of pure water at one atmospheric pressure (= 760 mm of Hg) is 100°C.

In a pressure cooker, steam is not allowed to escape out. The vapour pressure on water inside the pressure cooker becomes nearly 1·75 times the atmospheric pressure, so water boils in it at about 120°C to 125°C due to increase in pressure. Thus cooking of vegetables, etc. becomes much easier and faster in it since they get sufficient heat energy before the water boils. Further the cooked material inside it being at a higher temperature, remains warm over a long period.

At high altitudes, such as hills and mountains, the atmospheric pressure is low (less than one atmospheric pressure), therefore water boils at a temperature lower than 100°C and so it does not provide the required heat energy to its contents for cooking. Thus cooking there becomes very difficult and it takes a much longer time.

11.22 EFFECT OF IMPURITIES ON THE BOILING POINT

The boiling point of a liquid increases by addition of impurities to it. If common salt is added to water, it boils at a temperature higher than 100°C. Hence, we add salt while cooking pulses, because the salt in water provides sufficient heat energy to its contents. Cooking thus becomes easier and faster.

11.23 LATENT HEAT AND SPECIFIC LATENT HEAT

We have read that during the change of phase of a substance which takes place at a constant temperature, a considerable amount of heat energy is absorbed or liberated. *Heat energy is absorbed by a solid during melting and an equal amount of heat energy is liberated by the liquid during freezing, without showing any change in temperature.* Similarly, *heat energy is absorbed by a liquid during vaporisation and an equal amount of heat energy is liberated by the vapour during condensation, without showing any change in temperature.*

Since the heat energy absorbed (or librated) *in change of phase* is not externally manifested by any rise or fall in temperature, it is called the ***latent heat***. *Latent heat when expressed for unit mass of substance, is called the **specific latent heat**.*

Specific latent heat is denoted by the symbol L. Thus, specific latent heat

$$L = \frac{\text{heat absorbed (or librated) for the change of phase}}{\text{mass}}$$

If amount of heat energy Q is absorbed (or librated) by the mass m of a substance during its change of phase at a constant temperature, then specific latent heat

$$L = \frac{Q}{m} \qquad ...(11.15)$$

> Thus *specific latent heat of a phase is the quantity of heat energy absorbed (or librated) by the unit mass of the substance for the change in its phase at a constant temeprature.*

The amount of heat energy absorbed (or librated) by a given amount of substance for the change of phase for which specific latent heat is L, is given as :

$$Q = \text{mass } (m) \times \text{specific latent heat } (L) \qquad ..(11.16)$$

Unit of specific latent heat : The S.I. unit of specific latent heat is J kg^{-1}. Other common units are cal g^{-1} and kcal kg^{-1}. They are related as :

$$1 \text{ kcal kg}^{-1} = 1 \text{ cal g}^{-1}$$
$$1 \text{ cal g}^{-1} = 4{\cdot}2 \text{ J g}^{-1}$$
$$\text{or} \quad 1 \text{ cal g}^{-1} = 4{\cdot}2 \times 10^3 \text{ J kg}^{-1}$$

11.24 SPECIFIC LATENT HEAT OF FUSION OF ICE

> *The specific latent heat of melting of ice is the heat energy required to melt unit mass of ice at 0°Cto water at 0°C without any change in temperature.*
>
> On the other hand, *the specific latent heat of fusion of ice is the heat energy released when a unit mass of water at 0°C freezes to ice at 0°C without any change in temperature.*

Note : **(1)** For a pure substance, the specific latent heat of fusion is same as the specific latent heat of melting.

(2) For ice, the specific latent heat of fusion is 336000 J kg^{-1} (= 80 cal g^{-1}). It means that 1 kg of ice at 0°C absorbs 336000 J of heat energy to convert into water at 0°C (or 1 g ice at 0°C absorbs 80 cal of heat energy to convert into water at 0°C.)

In other words, 1 kg of water at 0°C will liberate 336000 J of heat energy to convert into ice at 0°C (or 1 g of water will liberate 80 cal of heat energy to convert into ice at 0°C).

(3) 1 g of water at 0°C has 336 J (or 80 cal) heat energy more than 1 g of ice at 0°C.

Specific latent heat of fusion of some common substances

Substance	*Latent heat of fusion*	
	cal g^{-1}	J kg^{-1}
Mercury	3	12·5 × 10^3
Sulphur	9	37·7 × 10^3
Silver	21	88·0 × 10^3
Paraffin wax	35	146 × 10^3
Copper	43	180 × 10^3
Ice	80	336 × 10^3

11.25 EXPLANATION OF LATENT HEAT OF MELTING ON THE BASIS OF KINETIC MODEL

According to the kinetic model, the molecules in a solid vibrate about their mean positions. The total energy of a molecule is the sum of its kinetic energy due to its motion (which depends on the temperature) and its potential energy (which depends on the force of attraction between the molecules and the separation between them). *When a solid changes into a liquid, without any change in temperature, the average kinetic energy of the molecules does not change, but the separation between the molecules on an average increases*. Some energy is required for increasing the separation against the attractive forces between the molecules (*i.e.*, for the increase in the potential energy of molecules). Since the increase in volume is negligible, so no heat is used in volume expansion. Thus the heat supplied during melting is utilised only in increasing the potential energy of the molecules and is called the latent heat of melting.

Note : When 1 g of ice at 0°C changes to water at 0°C, nearly 336 J energy is used in increasing the potential energy due to increase in intermolecular separation, so latent heat of ice is 336 J g^{-1}.

11.26 NATURAL CONSEQUENCES OF HIGH SPECIFIC LATENT HEAT OF FUSION OF ICE

(1) Snow on mountains does not melt all at once : The reason is that ice has a high specific latent heat of fusion (equal to 336000 J kg^{-1}). It is due to this fact that it changes into water slowly as it gets heat energy from the sun. If latent heat would have been low, all the snow would have melted in a short time on getting heat from the sun and there would have been flood in the rivers.

(2) In cold countries water in lakes and ponds does not freeze all at once : The reason is that the specific latent heat of fusion of ice is sufficiently high (= 336 J g^{-1}). The water in lakes and ponds will have to liberate a large quantity of heat to the surrounding before freezing. The layer of ice if formed on water, being a poor conductor of heat, will also prevent the loss of heat from water of lake, hence it does not freeze all at one.

(3) Drinks get cooled more quickly by adding pieces of ice than the ice-cold water at 0°C : This is because 1 g of ice at 0°C takes 336 J of heat energy from the drink to melt into water at 0°C. Thus the drink librates an additional 336 J of heat energy to 1 g ice at 0°C than to 1 g ice-cold water at 0°C. Therefore cooling produced by 1 g ice at 0°C is much more than that by 1 g water at 0°C.

(4) When ice in a frozen lake starts melting, its surrounding becomes very cold : The reason is that quite a large amount of heat energy is required for melting the frozen lake which is absorbed from the surrounding atmosphere. As a result, the temperature of the surrounding falls and it becomes very cold.

(5) It is generally more cold after a hail-storm (when ice melts) than during or before the hail-storm : The reason is that after the hail-storm, ice absorbs the heat energy required for its melting from the surroundings, so the temperature of the surroundings falls further down and we feel more cold.

EXAMPLES

1. A slab of ice at 0°C is constantly heated till the steam is formed at 100°C. Draw a graph showing the change in temperature with time. Label the various parts of the graph properly.

The graph showing the change of temperature with time is given in Fig. 11.5.

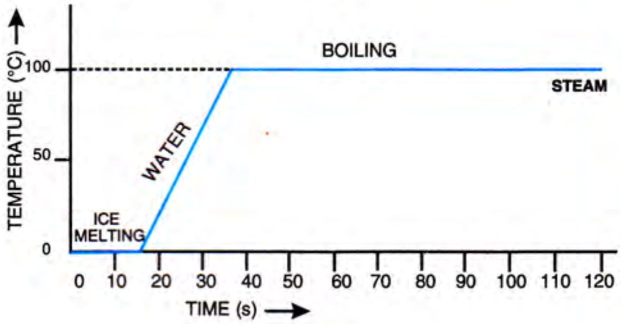

Fig. 11.5

2. A piece of ice is heated at a constant rate. The variation in temperature with time of heating is shown in the graph in Fig. 11.6.

(i) What is represented by the part AB ?

(ii) What does the part CD represent ?

(iii) What conclusion do you draw regarding the nature of ice from the graph ?

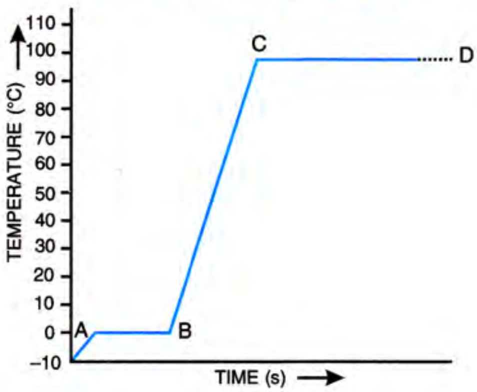

Fig. 11.6

(i) The part AB represents melting of ice at 0°C.

(ii) The part CD represents boiling of water at 100°C.

(iii) The ice initially is in solid state at –10°C. On heating, its temperature rises to 0°C. It then takes some heat energy at 0°C to melt in water at 0°C which is called its latent heat.

3. A substance initially in solid state at 0°C is heated. The graph showing the variation in temperature with the amount of heat supplied is shown in Fig. 11.7. If the specific heat capacity of the solid substance is 500 J kg^{-1} $°C^{-1}$, use graph to find : (i) the mass of the substance, and (ii) the specific latent heat of fusion of the substance in the liquid state.

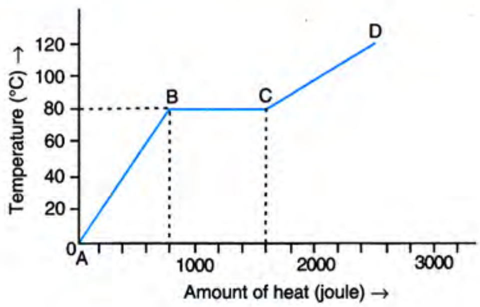

Fig. 11.7

(i) In the graph, part AB represents the increase in temperature of solid substance from 0°C to 80°C on absorbing the heat Q = 800 J. If m kg is the mass of solid substance, then

Heat absorbed Q = mass × specific heat capacity × rise in temperature

or $800 = m \times 500 \times (80 - 0)$

$\therefore \quad m = \frac{800}{500 \times 80} = \frac{1}{50}$ kg

= **0·02 kg** (or **20 g**).

(ii) In the graph, part BC represents the change of state of the substance from solid to liquid at 80°C on absorbing the heat $Q = 1600\text{ J} - 800\text{ J} = 800\text{ J}$. If $L\text{ J kg}^{-1}$ is the latent heat of fusion of the substance in the liquid state, then

Heat absorbed Q = mass × latent heat of fusion

or $800 = 0{\cdot}02 \times L$

$\therefore\ L = \dfrac{800\text{ J}}{0{\cdot}02\text{ kg}} = $ **40000 J kg^{-1}.**

4. How much heat energy is required to melt 5 kg of ice ? Specific latent heat of ice = 336 J g^{-1}.

Given, $m = 5\text{ kg} = 5000\text{ g}$, $L = 336\text{ J g}^{-1}$

Heat energy required $= mL$

$= 5000\text{ g} \times 336\text{ J g}^{-1}$

$= 1680000\text{ J} =$ **$1{\cdot}68 \times 10^6$ J**

5. The temperature of 300 g of water at 40°C is lowered to 0°C by adding ice to it. Find the mass of ice added if specific heat capacity of water is 4·2 J g^{-1} K^{-1} and specific latent heat of ice is 336 J g^{-1}.

Given : mass of water $m = 300$ g, initial temperature = 400°C, final temperature = 0°C

$\therefore$ Fall in temperature $\Delta t = (40 - 0) = 40°\text{C} = 40\text{ K}$

Heat lost by water $= m\,c\,\Delta t$

$= 300 \times 4{\cdot}2 \times 40$

$= 5{\cdot}04 \times 10^4\text{ J}$ ...(i)

If m'g ice is added, heat gained by it to melt to 0°C

$= m'L$

$= m' \times 336\text{ J}$ ...(ii)

By the principle of method of mixture,

heat lost by water = heat gained by ice

$\therefore\ 5{\cdot}04 \times 10^4 = m' \times 336$

or $m' = \dfrac{5{\cdot}04 \times 10^4}{336} =$ **150 g**

6. How much boiling water at 100°C is needed to melt 2 kg of ice so that the mixture, which is all water, is at 0°C ? Given : specific heat capacity of water = 4·2 J g^{-1} K^{-1}, specific latent heat of ice = 336 J g^{-1}.

Given, mass of ice = 2 kg = 2000 g

Let m g of boiling water be required.

Assuming that no heat energy is lost,

heat energy given by boiling water

= heat energy required for ice to melt

or mass of boiling water × specific heat capacity of water × fall in temperature

= mass of ice × specific latent heat of ice

or $m \times 4{\cdot}2 \times (100 - 0) = 2000 \times 336$

or $m = \dfrac{2000 \times 336}{4{\cdot}2 \times 100} = 1600\text{ g} =$ **1·6 kg.**

7. Heat energy is supplied at a constant rate to 400 g of ice at 0°C. The ice is converted into water at 0°C in 5 minutes. How much time will be required to raise the temperature of water from 0°C to 100°C ?

Specific latent heat of ice = 336 J g^{-1}, specific heat capacity of water = 4·2 J g^{-1} K^{-1}.

Heat energy required to melt 400 g of ice at 0°C

$= mL = 400 \times 336 = 134{,}400\text{ J}$

This heat energy is supplied in 5 minutes.

$\therefore$ Heat energy supplied per minute $= \dfrac{134{,}400\text{ J}}{5\text{ min}}$

$= 26880\text{ J min}^{-1}$

Heat energy required to raise the temperature of water from 0 °C to 100 °C

$= m \times c \times$ rise in temperature

$= 400 \times 4{\cdot}2 \times 100 = 168000\text{ J}$ ...(i)

If time required is t minute, then

heat energy supplied in t minutes $= 26880 \times t$ J ..(ii)

From eqns, (i) and (ii),

$26880 \times t = 168000$

or time required $t = \dfrac{168000}{26880}$

$= 6{\cdot}25$ minute = **6 min and 15 s**

8. Calculate the power of an electric heater required to melt 1 kg of ice at 0°C in 30 s if the efficiency of heater is 40%. Take specific latent heat of ice = 336 J g^{-1}.

Given m = 1kg, efficiency = 40%, t = 30 s, $L = 336\text{ J g}^{-1} = 336 \times 10^3\text{ J kg}^{-1}$

If P watt is the power of heater, then energy supplied by the heater in time t s $= Pt$ joule

Since the efficiency of heater is 40%, so 40% of the energy supplied by heater is used in melting the ice.

i.e., $\dfrac{40}{100} Pt = mL$

or $0{\cdot}4\,P \times 30 = 1 \times 336 \times 10^3$

or $P = \dfrac{336 \times 10^3}{0{\cdot}4 \times 30}$

$= 28 \times 10^3$ watt or **28 kW**

9. One kilogram of ice at –10°C is heated at a constant rate until the whole of it vaporises. How much heat is required ? Specific latent heat of fusion of ice = 336×10^3 J kg^{-1}, specific latent heat of steam = 2268×10^3 J kg^{-1}, specific heat capacity of ice = $2{\cdot}1 \times 10^3$ J kg^{-1} K^{-1}, specific heat capacity of water = $4{\cdot}2 \times 10^3$ J kg^{-1} K^{-1}.

Given, m = 1 kg, initial temperature = –10°C

In this case, heat energy is required in the following *four* steps :

(i) Heat energy required to raise the temperature of ice from –10°C to 0°C

$$Q_1 = m \times c_{ice} \times \text{rise in temperature}$$
$$= 1 \times (2{\cdot}1 \times 10^3) \times [0 - (-10)] = 21 \times 10^3 \text{ J}$$

(ii) Heat energy required to melt the ice at 0°C into water at 0°C,

$$Q_2 = mL_{ice} = 1 \times (336 \times 10^3)$$
$$= 336 \times 10^3 \text{ J}$$

(iii) Heat energy required to increase the temperature of melted ice-water from 0°C to 100°C

$$Q_3 = m \times c_{water} \times \text{rise in temperature}$$
$$= 1 \times (4{\cdot}2 \times 10^3) \times (100 - 0)$$
$$= 420 \times 10^3 \text{ J}$$

(iv) Heat energy required for vaporisation of water at 100°C into steam at 100°C

$$Q_4 = mL_{steam} = 1 \times (2268 \times 10^3)$$
$$= 2268 \times 10^3 \text{ J}$$

∴ Total heat energy required = $Q_1 + Q_2 + Q_3 + Q_4$

$$= (21 + 336 + 420 + 2268) \times 10^3$$
$$= 3045 \times 10^3 \text{ J} = \mathbf{3045 \text{ kJ}}$$

Note : In this case, on the temperature-time graph, the time in four steps will be proportional to the amount of heat energy taken in these steps *i.e.*, it will be in the ratio 21 : 336 : 420 : 2268 or 1 : 16 : 20 : 108 as heat is being supplied at a constant rate.

10. A cube of ice of mass 30 g at 0°C is added into 200 g of water at 30°C. Calculate the final temperature of water when whole of the ice cube has melted.

Given: specific latent heat of ice = 80 cal g^{-1}, specific heat capacity of water = 1 cal g^{-1} °C^{-1}.

Let final temperature of water be t°C.

Heat energy taken by the ice to melt at 0°C

$$Q_1 = mL$$
$$= 30 \text{ g} \times 80 \text{ cal g}^{-1} = 2400 \text{ cal}$$

Heat energy taken by the melted ice to raise its temperature from 0°C to t°C

$$Q_2 = \text{mass} \times \text{specific heat capacity} \times \text{rise in temperature}$$
$$= 30 \text{ g} \times 1 \text{ cal g}^{-1} \text{ °C}^{-1} \times t \text{ °C}$$
$$= 30\,t \text{ cal}$$

Total heat energy taken by ice = $Q_1 + Q_2$

$$= (2400 + 30\,t) \text{ cal} \quad ....(i)$$

Heat energy given by water in fall of its temperature from 30 °C to t °C

$$= \text{mass} \times \text{specific heat capacity} \times \text{fall in temperature}$$
$$= 200 \text{ g} \times 1 \text{ cal g}^{-1} \text{ °C}^{-1} \times (30 - t)\text{°C}$$
$$= 200\,(30 - t) \text{ cal.} \quad ....(ii)$$

If there is no heat loss,

heat energy given by water

= total heat energy taken by ice

$$200\,(30 - t) = 2400 + 30\,t$$

or $6000 - 200\,t = 2400 + 30\,t$

or $230\,t = 3600$

or $t = \frac{3600}{230} = \mathbf{15{\cdot}65 \text{ °C}}$

11. A vessel of mass 100 g contains 150 g of water at 30°C. How much ice is needed to cool it to 5 °C ? Take specific heat capacity of material of vessel = $0{\cdot}4$ J g^{-1} K^{-1}, specific latent heat of fusion of ice = 336 J g^{-1} and specific heat capacity of water = $4{\cdot}2$ J g^{-1} K^{-1}.

Heat energy imparted by vessel and water contained in it in cooling from 30°C to 5°C, is used in melting ice and then raising the temperature of melted ice from 0°C to 5°C.

Heat energy imparted by vessel

$$Q_1 = 100 \times 0{\cdot}4 \times (30 - 5) = 1000 \text{ J} \quad ....(i)$$

Heat energy imparted by water

$$Q_2 = 150 \times 4{\cdot}2 \times (30 - 5) = 15750 \text{ J} \quad ....(ii)$$

Let m g of ice be used, then

Heat energy taken by ice to melt

$$Q_3 = m \times 336 \text{ J} \quad ....(iii)$$

and heat energy taken by the melted ice to raise its temperature from 0°C to 5°C

$$Q_4 = m \times 4{\cdot}2 \times (5 - 0) = 21\,m \text{ J} \quad ....(iv)$$

By law of conservation of energy,

heat energy imparted by vessel and water

= heat energy taken by ice and melted ice or $Q_1 + Q_2 = Q_3 + Q_4$

i.e., $1000 + 15750 = 336\,m + 21\,m$

or $357\,m = 16750$

or $m = \frac{16750}{357} = 46{\cdot}92$ g

Thus, **46·92 g** of ice is used.

12. A vessel of negligible heat capacity contains 5·0 kg of water at 50°C. If 5·0 kg of ice at 0°C is added to it, find : (i) heat energy imparted by water in fall of its temperature from 50°C to 0°C, (ii) mass of ice melted, (iii) final temperature of mixture, and (iv) mass of water at 0°C in mixture. Given : specific heat capacity of water = 4200 J kg^{-1} K^{-1}, specific latent heat of ice = 336 kJ kg^{-1}.

(i) Heat energy imparted by water in fall of its temperature from 50°C to 0°C

= mass × specific heat capacity × fall in temperature

$= 5{\cdot}0 \times 4200 \times (50 - 0)$ J = **1050000 J**

(ii) If m kg of ice melts at 0°C, heat energy used in melting = $m\,L = m \times (336 \times 10^3)$ J

If there is no loss of energy,

heat energy imparted = heat energy used in melting

or $1050000 \text{ J} = m \times (336 \times 10^3)$ J

$\therefore$ $m = \frac{1050000}{336000} = $ **3·125 kg**

Remaining ice = 5·0 – 3·125 = **1·875 kg.**

Thus **1·875** kg ice remains unmelted at 0°C.

(iii) Final temperature of mixture will be **0°C** [since the mixture will contain 5·0 kg of water at 0°C + 3·125 kg ice-melted water at 0°C + 1·875 kg unmelted ice at 0°C].

(iv) Mass of water at 0°C in mixture = 5·0 + 3·125

= **8·125 kg.**

13. What will be the result of mixing 400 g of copper chips at 500°C with 500 g of crushed ice at 0°C ?

Specific heat capacity of copper = 0·42 J g^{-1} K^{-1}, specific latent heat of fusion of ice = 340 J g^{-1}.

Given, mass of copper chips = 400 g,

temperature of copper chips = 500°C,

mass of ice = 500 g

Heat energy imparted by copper chips in cooling from 500°C to 0°C

$Q_1 = 400 \text{ g} \times 0{\cdot}42 \text{ J g}^{-1} \text{ K}^{-1} \times (500 - 0) \text{ K} = 84000$ J

Heat energy required to melt 500 g ice at 0°C

$Q_2 = 500 \times 340 = 170000$ J

Since $Q_2 > Q_1$, therefore all ice will not melt.

Let m g of ice melts, then

$m \times 340 = 84000$

$\therefore$ $m = \frac{84000}{340} = $ **247 g**

Thus only **247 g** ice will melt and the temperature will remain **0°C**.

14. An electric heater of power 150 W is immersed in 0·75 kg of ice at 0°C in a lagged container of negligible heat capacity. The temperature remains constant for 27·5 minutes and then rises to 40·0°C in a further 14 minutes. Explain why does the temperature remain constant. Calculate : (a) the specific latent heat of ice, and (b) the specific heat capacity of water.

Heat energy supplied by heater is used for melting the ice in the first 27·5 minutes and therefore the temperature remains constant during this time.

Given : power of heater = 150 W = 150 J s^{-1}.

i.e., heater supplies energy 150 joule per second.

(a) Heat energy supplied by the heater in 27·5 minute (or 27·5 × 60 s) = power × time

$= 150 \times (27{\cdot}5 \times 60)$ J = 247500 J ...(i)

Let L J kg^{-1} be the specific latent heat of ice.

Heat energy used in melting the ice = mL

$= 0{\cdot}75\,L$ ...(ii)

Now heat energy used in melting

= heat energy supplied

i.e., $0{\cdot}75 \times L = 247500$

or $L = \frac{247500}{0.75} = $ **330×10^3 J kg^{-1}**

(b) Heat energy supplied by heater in 14 minutes (or 14 × 60 s) = power × time

$= 150 \times (14 \times 60) = 126{,}000$ J ...(i)

Let c J kg^{-1} K^{-1} be the specific heat capacity of water.

Heat energy taken by water at 0°C to raise it to 40°C = $m \times c \times$ rise in temperature

$= 0{\cdot}75 \times c \times (40 - 0)$ J $= 30\,c$ J ...(ii)

Now heat energy taken by water

= heat energy supplied by heater

$\therefore$ $30\,c = 126{,}000$

or $c = \frac{126000}{30} = $ **$4{\cdot}2 \times 10^3$ J kg^{-1} K^{-1}**

EXERCISE–11(B)

1. (a) What do you understand by the change of phase of a substance ?
   (b) Is there any change in temperature during the change of phase ?
   (c) Does the substance absorb or liberate any heat energy during the change of phase ?
2. Explain the terms melting and melting point.
3. A substance on heating, undergoes (i) a rise in its temperature, (ii) a change in its phase without change in its temperature. In each case, state the change in energy of molecules of the substance.

Ans. (i) averge kinetic energy of molecules increases, (ii) average potential energy of molecules increases

4. How does the (a) average kinetic energy (b) average potential energy of molecules of a substance change during its change in phase at a constant temperature, on heating ?

Ans. (a) it does not change, (b) it increases

5. State the effect of presence of impurity on the melting point of ice. Give *one* use of it.
6. State the effect of increase of pressure on the melting point of ice.
7. The diagram in Fig. 11.8 below shows the change of phases of a substance on a temperature-time graph on heating the substances at a constant rate.

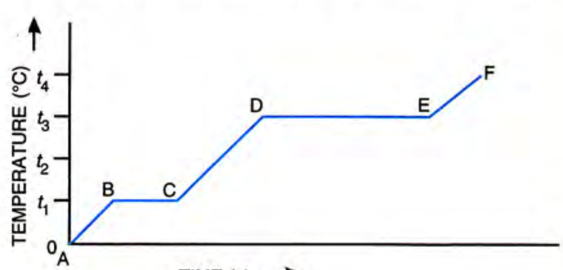

Fig. 11.8

   (a) What do the parts AB, BC, CD and DE represent ?
   (b) What is the melting point of the substance ?
   (c) What is the boiling point of the substance ?

Ans. (a) AB part – rise in temperature of solid from 0°C to t_1°C, BC part – melting at temperature t_1°C, CD part – rise in temperature of liquid from t_1°C to t_3°C, DE part – boiling at temperature t_3°C (b) t_1°C (c) t_3°C.

8. 1 kg of ice at 0°C is heated at a constant rate and its temperature is recorded after every 30 s till steam is formed at 100°C. Draw a temperature-time graph to represent the change of phases.
9. Explain the terms boiling and boiling point.
10. How is the volume of water affected when it boils at 100°C ?
11. How is the boiling point of water affected when some salt is added to it ?
12. What is the effect of increase in pressure on the boiling point of a liquid ?
13. Water boils at 120°C in a pressure cooker. Explain the reason.
14. Write down the approximate range of temperature at which water boils in a pressure cooker.

Ans. 120°C to 125°C

15. It is difficult to cook vegetables on hills and mountains. Explain the reason.
16. Complete the following sentences :
   (a) When ice melts, its volume ....
   (b) Decrease in pressure over ice .... its melting point.
   (c) Increase in pressure .... the boiling point of water.
   (d) A pressure cooker is based on the principle that boiling point of water increases with the ....
   (e) The boiling point of water is defined as ....
   (f) Water can be made to boil at 115°C by .... pressure over its surface.

Ans. (a) decreases (b) increases, (c) increases, (d) increase in pressure, (e) the constant temperature at which water changes to steam, (f) increasing.

17. What do you understand by the term latent heat ?
18. Define the term specific latent heat of fusion of ice. State its S.I. unit.
19. Write the approximate value of specific latent heat of ice.

Ans. 336 kJ kg^{-1}

20. 'The specific latent heat of fusion of ice is 336 J g^{-1}'. Explain the meaning of this statement.

21. 1 g ice at 0°C melts to form 1 g water at 0°C. State whether the latent heat is absorbed or given out by ice. **Ans.** latent heat is absorbed by ice.

22. Which has more heat : 1 g of ice at 0°C or 1 g of water at 0°C ? Give reason.

Ans. 1 g of water at 0°C

Reason : 1 g of water at 0°C liberates 80 cal heat to form 1 g of ice at 0°C.

23. (a) Which requires more heat : 1 g ice at 0°C or 1 g water at 0°C to raise its temperature to 10°C ?

(b) Explain your answer in part (a).

Ans. (a) 1 g ice at 0°C

(b) 1 g ice at 0°C first absorbs 336 J heat to convert into 1 g water at 0°C.

24. Ice cream appears colder to the mouth than water at 0°C. Give reason.

25. The soft drink bottles are cooled by (i) ice cubes at 0°C, and (ii) iced-water at 0°C. Which will cool the drink quickly ? Give reason.

26. It is generally cold after a hail-storm than during and before the hail-storm. Give reason.

27. The temperature of the surrounding starts falling when ice in a frozen lake starts melting. Give reason.

28. Water in lakes and ponds do not freeze at once in cold countries. Give reason.

29. Explain the following :

(a) The surrounding become pleasantly warm when water in a lake starts freezing in cold countries.

(b) The heat supplied to a substance during its change of state, does not cause any rise in its temperature.

MULTIPLE CHOICE TYPE

1. The S.I. unit of specific latent heat is :

(a) cal g^{-1} (b) cal g^{-1} K^{-1}

(c) J kg^{-1} (d) J kg^{-1} K^{-1} **Ans.** (c) J kg^{-1}

2. The specific latent heat of fusion of water is :

(a) 80 cal g^{-1} (b) 2260 J g^{-1}

(c) 80 J g^{-1} (d) 336 J kg^{-1}

Ans. (a) 80 cal g^{-1}

NUMERICALS

1. 10 g of ice at 0°C absorbs 5460 J of heat energy to melt and change to water at 50°C. Calculate the specific latent heat of fusion of ice. Specific heat capacity of water is 4200 J kg^{-1} K^{-1}.

Ans. 336 J g^{-1}

2. How much heat energy is released when 5·0 g of water at 20°C changes into ice at 0°C ? Take specific heat capacity of water = 4·2 J g^{-1} K^{-1}, specific latent heat of fusion of ice = 336 J g^{-1}.

Ans. 2100 J

3. A molten metal of mass 150 g is kept at its melting point 800°C. When it is allowed to freeze at the same temperature, it gives out 75,000 J of heat energy.

(a) What is the specific latent heat of the metal ?

(b) If the specific heat capacity of metal is 200 J kg^{-1} K^{-1}, how much additional heat energy will the metal give out in cooling to – 50 °C ?

Ans. (a) 500 J g^{-1}, (b) 25,500 J

4. A refrigerator converts 100 g of water at 20°C to ice at – 10°C in 73·5 min. Calculate the average rate of heat extraction in watt. The specific heat capacity of water is 4·2 J g^{-1} K^{-1}, specific latent heat of ice is 336 J g^{-1} and the specific heat capacity of ice is 2·1 J g^{-1} K^{-1}. **Ans.** 10 W

5. In an experiment, 17 g of ice is used to bring down the temperature of 40 g of water at 34°C to its freezing temperature. The specific heat capacity of water is 4·2 J g^{-1} K^{-1}. Calculate the specific latent heat of ice. State *one* important assumption made in the above calculation. **Ans.** 336 J g^{-1}

Assumption : No loss of energy.

6. Find the result of mixing 10 g of ice at –10°C with 10 g of water at 10°C. Specific heat capacity of ice = 2·1 J g^{-1} K^{-1}, specific latent heat of ice = 336 J g^{-1} and specific heat capacity of water = 4·2 J g^{-1} K^{-1}.

Ans. 0·625 g ice will melt and temperature will remain 0°C

7. A piece of ice of mass 40 g is added to 200 g of water at 50°C. Calculate the final temperature of water when all the ice has melted. Specific heat capacity of water = 4200 J kg^{-1} K^{-1} and specific latent heat of fusion of ice = 336 × 10^3 J kg^{-1}.

Ans. 28·33°C

8. 250 g of water at 30°C is contained in a copper vessel of mass 50 g. Calculate the mass of ice required to bring down the temperature of the vessel

and its contents to 5°C. Given : specific latent heat of fusion of ice = 336×10^3 J kg^{-1}, specific heat capacity of copper = 400 J kg^{-1} K^{-1}, specific heat capacity of water = 4200 J kg^{-1} K^{-1}.

Ans. 74·93 g

9. 2 kg of ice melts when water at 100°C is poured in a hole drilled in a block of ice. What mass of water was used ? Given : specific heat capacity of water = 4200 J kg^{-1} K^{-1}, specific latent heat of ice = 336×10^3 J kg^{-1}. **Ans.** 1·6 kg

10. Calculate the total amount of heat energy required to convert 100 g of ice at –10°C completely into water at 100°C. Specific heat capacity of ice = 2·1 J g^{-1} K^{-1}, specific heat capacity of water = 4·2 J g^{-1} K^{-1}, specific latent heat of ice = 336 J g^{-1}.

Ans. $7{\cdot}77 \times 10^4$ J

11. The amount of heat energy required to convert 1 kg of ice at –10°C to water at 100°C is 7,77,000 J. Calculate the specific latent heat of ice. Specific heat capacity of ice = 2100 J kg^{-1} K^{-1}, specific heat capacity of water = 4200 J kg^{-1} K^{-1}.

Ans. 3,36,000 J kg^{-1}

12. 200 g of ice at 0°C converts into water at 0°C in 1 minute when heat is supplied to it at a constant rate. In how much time 200 g of water at 0°C will change to 20°C ? Take specific latent heat of ice = 336 J g^{-1}. **Ans.** 15 s