UNIT 2 HEAT

Structure

2.1 INTRODUCTION

In the previous unit, you have learnt certain concepts in mechanics. In this unit, we discuss the concepts of heat, a form of energy. You will learn about the most common effect of heat on fluids, i.e. real and apparent expansions of fluids. We will describe simple experiments to demonstrate these effects and define different coefficients with their relationships. We will explain the anomalous behaviour of water with Hope's apparatus and the volume v/s temperature graph.

You will study the method of measuring heat using appropriate instruments or devices. We will also define thermal properties like specific heat, latent heat and describe the methods of determining these quantities. We will discuss the effects of the physical parameters like pressure on melting and boiling points.

You must have noticed the ways of heat transmissions around you – the conduction, convection and radiation. In solids, heat is transmitted by conduction, in fluids by convection and in vacuum by radiation. You will study about these in some detail.

Throughout the unit and at the end we give some problems which helps you for better understanding the theory and clarifying some concepts. These problems are also intended in check your progress.

Objectives

After studying this unit, you should be able to

- describe the real and apparent expansion of liquid,
- define the coefficients of expansion of liquid and derive the relations between them,
- discuss the anomalous behaviour of water using Hope's experiment,
- define specific heat and latent heat, and describe the methods for determining them,
- discuss the effect of physical parameters on melting and boiling points, and
- explain the ways in which heat is transmitted.

2.2 THERMAL EXPANSION OF LIQUID

Real and Apparent Expansions of a Liquid and their Relationship

You may know that substance which is capable of flowing is called a **fluid**. Both liquid and gases are fluids. They do not have any definite length and surface area. What happens when we heat them?

On heating liquids, they can undergo cubical expansion only, i.e. they undergo change in volume only. But when we heat a liquid, the container also expands. Thus, we get expansion of both the container and liquid.

If we measure the increase in the volume of a liquid without considering the expansion of the container, then the observed increase in volume of the liquid is called the **apparent expansion** of the liquid. On the other hand, if we take into consideration the expansion of the container also, then the increase in volume is the real expansion of the liquid.

The real expansion of a liquid is always greater than its apparent expansion. We can show this by the following experiment.

Relation between Real and Apparent Expansion, using a Dilatometer

This apparatus consists of a glass bulb having a long and narrow stem (Figure 2.1). A scale marked on the stem reads the volume of the dilatometer up to each mark.

We fill the dilatometer up to the mark A and place it on a heater so that its temperature can rise by 10°C. It is observed that initially the level of liquid falls to the mark B and then rises to the mark C.

This is because initially only the container gets heated and its volume increases. Hence, there is a drop in the level of the liquid. Thereafter, the liquid also begins to expand and its level goes from B to C. The expansion from A to C is the apparent expansion of the liquid. (AB represents expansion of the container.)

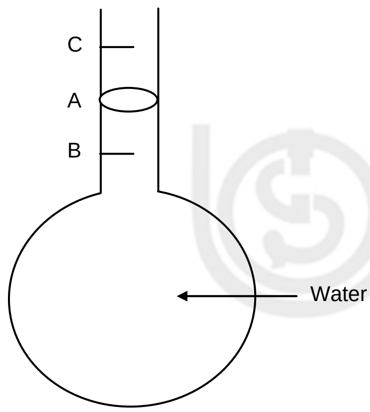

Figure 2.1 : Demonstrating Real and Apparent Expansion

Let V_0 be the volume of the liquid at 0°C. Let the liquid be heated to t°C.

$$BC = AB + AC$$

where, BC = Real Expansion of liquid,

AC = Apparent Expansion of liquid, and

AB = Cubical Expansion of container.

The coefficient of real expansion of liquid (γ_r) is defined as the real increase in the volume per unit original volume at 0°C per degree rise in temperature.

$$\gamma_r = \frac{BC}{V_0\, t} \quad \ldots (2.1)$$

The coefficient of apparent expansion of liquid γ_a is defined as the apparent increase per unit original volume at 0°C per degree rise in temperature. Then

$$\gamma_a = \frac{AC}{V_0\, t} \quad \ldots (2.2)$$

Relation between Two Coefficients

We have $$\gamma_r = \frac{AC + AB}{V_0\, t}$$

$$= \frac{AC}{V_0\, t} + \frac{AB}{V_0\, t} = \gamma_a + \frac{AB}{V_0\, t} \cdot \frac{V_C}{V_C}$$

$$= \gamma_a + \frac{AB}{V_C\, t} \frac{V_C}{V_0} = \gamma_a + \gamma_C \frac{V_C}{V_0} \qquad \ldots (2.3)$$

where, V_C is the volume of container at 0°C and γ_C is the coefficient of expansion of the material of the container.

Different liquids have different coefficients of expansion and we can understand this by the following experiment.

Relative Expansion of Different Liquids

A number of liquids are taken up to the same level in identical flasks. The volumes of the liquids are the same. The flasks are put in a common water bath in which the constant temperature is maintained above room temperature.

Now the levels of the liquids are observed. They are found to be different. This shows that for the same rise in temperature the different liquids expand by different amounts, i.e. different liquids have different coefficients of expansions.

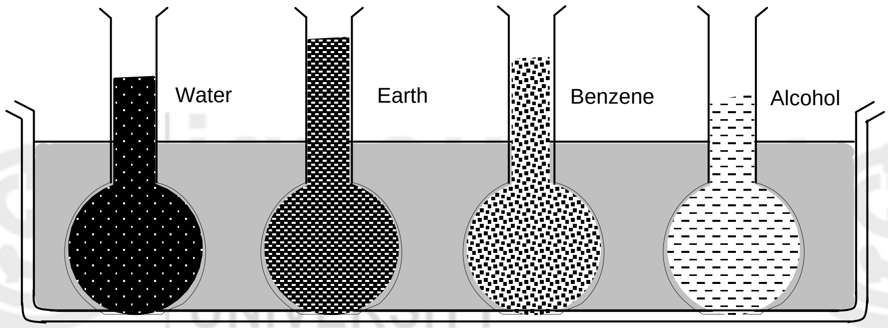

Figure 2.2 : Relative Expansion of Different Liquids

Change in Density with Temperature

When the temperature of a given mass of a liquid changes; its mass remains the same but its volume changes. Therefore, the density also changes.

Let V_0 and V be the volumes of the liquid at 0°C and t°C respectively and let d_0 and d be the densities respectively, then you know that :

$$m = V_0\, d_0 \text{ and } m = Vd$$

Since $$V = V_0\, (1 + \gamma\, t)$$

$$\frac{m}{d} = \frac{m\,(1 + \gamma\, t)}{d_0}$$

$$d = \frac{d_0}{1 + \gamma\, t} = d_0\, (1 - \gamma\, t) \qquad \ldots (2.4)$$

Since γt is usually much less than 1.

Liquids generally expand on heating and contract on cooling. However, water is an outstanding exception between 0°C and 4°C. You should learn about the unusual expansion of water.

Anomalous (Unusual) Expansion of Water

If we go on cooling water then we find that its volume decreases till it reaches 4^{o}C but as it is cooled below 4^{o}C its volume increases instead of decreasing.

This expansion of water when it is cooled below 4^{o}C is known as the **anomalous expansion** of water.

Let us consider an experiment for the explanation using "**Hope's apparatus**" :

The apparatus consists of a tall metal cylinder provided with two openings for thermometers T_1 and T_2 (Figure 2.3). The central part of the apparatus is surrounded by a circular trough in which freezing mixture of ice and salt is present. The cylinder is filled with water of temperature 10^{o}C. Water is then allowed to cool down.

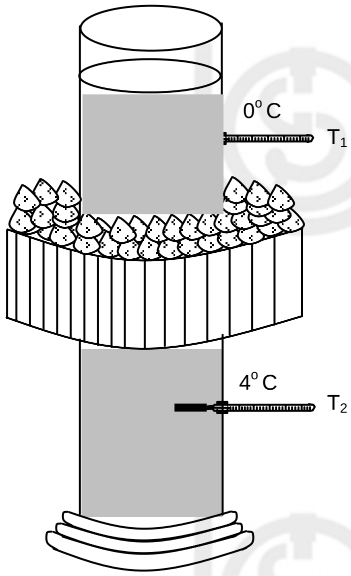

Figure 2.3 : Hope's Apparatus

(i) Initially both the thermometers show the same temperature. The temperature shown by the lower thermometer then slowly falls to 4^{o}C.

(ii) The temperature of the lower thermometer remains constant at 4^{o}C while the temperature of the upper thermometer slowly goes down until it reaches 0^{o}C.

(iii) After sometime a thin film of ice is formed at the surface of water in the cylinder.

The above observation can be explained in the following manner :

The temperature of water in the middle portion of the cylinder decreases due to the freezing mixture surrounding it. As it cools, its volume decreases and hence its density increases and it sinks to the bottom of the cylinder. Therefore, the temperature in the lower thermometer goes on decreasing upto 4^{o}C. As the water in the middle portion of the cylinder gets cooled below 4^{o}C, it begins to expand so that its density decreases and hence it cannot sink to the bottom. Therefore, water in the lower part of the cylinder remains at 4^{o}C . This constant temperature is now shown by the lower thermometer.

At this stage the temperature in the upper part of the cylinder begins to decrease, falls to 4^{o}C and continues to decrease upto 0^{o}C. This shows that as water cools below 4^{o}C, its density decreases and hence it rises to the surface.

As ice floats on the surface water, its density must be less than that of water. The temperature shown by the lower thermometer remains constant at 4^{o}C, while that shown by the upper thermometer falls to 0^{o}C. This shows that the density of water is maximum at 4^{o}C.

If we plot a graph showing the relation between temperature and the volume of 1 kg of water, then we get the curve shown in Figure 2.4.

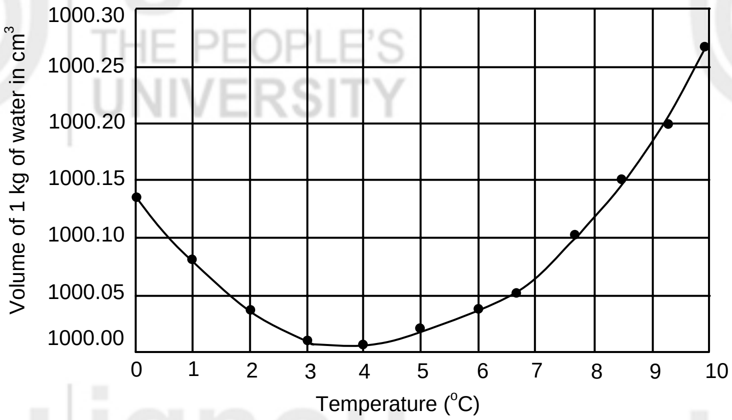

Figure 2.4 : Relation between Temperature and Volume of 1 kg Water

So far you have studied about the effect of heat on the expansion of fluids. You may like to know, how do we measure heat?

SAQ 1

(a) A glass flask with volume 200 cm^3 is filled to the brim with mercury at $20^{o}C$. How much mercury overflows when the temperature of the system is raised to $100^{o}C$? The coefficient of the volume expansion of glass is $1.2 \times 10^{-5}\ K^{-1}$ and the coefficient of volume expansion of mercury is $18 \times 10^{-5}\ K^{-1}$.

(b) A one litre flask contains some mercury. It is found that at different temperatures, the volume of air inside the flask remains the same. What is the volume of mercury in flask? Given that coefficient of cubical expansion of glass = $27 \times 10^{-6}\ C^{-1}$ and coefficient of volume expansion of $H_g = 1.8 \times 10^{-4}\ C^{-1}$.

(c) The coefficient of volume expansion of glycerine is $49 \times 10^{-5}\ {}^{o}C^{-1}$. What is the fractional change in the density for a $30^{o}C$ rise in temperature?

(d) The density of mercury is 13.6 g cm^{-3} at $0^{o}C$ and if its coefficient of cubical expansion is $1.82 \times 10^{-4}\ {}^{o}C^{-1}$, calculate the density of mercury at $50^{o}C$.

2.3 MEASUREMENT OF HEAT

You know that when we measure a physical quantity, we need to specify the units of measurement. So you must first learn about the units in which the heat content of a body is measured.

Units of Heat

The commonly used unit of heat is kilocalorie (kcal).

A kilocalorie is the amount of heat required to raise the temperature of 1 kilogram of water through 1°C (from 14.5°C to 15.5°C).

This range of temperatures is special in defining the unit quantity of heat energy, because water shows uniform and smooth expansion beyond 14°C and slightly different quantities of heat are required for raising the temperature through 1°C from different temperatures for 1 kg of water. However, since the difference in the heat required is small, *we can say that one kilocalorie is the quantity of heat required to raise the temperature of one kilogram of water through 1°C.*

In CGS system the unit of heat is calorie and is defined as the quantity of heat required to raise the temperature of one gram of water through 1°C.

1 kilocalorie = 1000 calories

Sometimes, heat is measured in mechanical or work units known as Joule (J), where we define

1 kcal = 4200 J

or 1 calorie = 4.18 J

Heat Capacity

The heat required to raise the temperature through 1°C depends upon the mass and also on the material of the substance. For example, if *equal masses* of water and copper are heated through 1°C, then a large amount of heat has to be supplied to water.

The quantity of heat required to raise the temperature through 1°C is called its heal capacity (or thermal capacity).

The units of heat capacity are : kilocalorie per °C (kcal /°C), calorie per °C (cal/°C) and joules per °C (J/°C).

The last one is the unit used in physics.

Specific Heat Capacity

As seen above, different quantities of heat are required to change the temperature of same mass of different materials through the same range of temperatures. This shows that the different materials must have a property by virtue of which the same mass when heated through the same range of temperatures would absorb different quantities of heat. This thermal property of the substance is called the *specific heat capacity* of the substance. It is also called the specific heat of the substance.

Specific heat capacity is defined as the quantity of heat energy required to raise the temperature of a unit mass of a substance through 1^oC. It is denoted by C.

Units of specific heat are : kilocalorie per kg per °C (kcal/kg°C), calorie per gm per °C (cal/gm°C), and joules per kg per ^oC (J/ kg°C).

Again, the last one is the unit used in physics.

For example, if the specific heat capacity of a substance is C, it means that one kg of that substance requires C joules to raise its temperature through 1^oC. Therefore, if m gram of the same substance is heated to raise its temperature through 1^oC, then the amount of heat required is 'mC' J. But by definition, this is the heat capacity of a substance of mass m.

Heat Capacity = mC

or Heat Capacity = Mass × Specific Heat Capacity

Since, the heat required to raise the temperature of mass m by 1°C is mC, the heat required (Q) to raise the temperature by θ^oC is given as

$$Q = mC\ .\ \theta$$

or Q = Heat Capacity × Rise in Temperature

or Q = Mass × Specific Heat Capacity ×Rise in Temperature

Let us apply this formula to solve a problem.

Example 2.1

Calculate the amount of heat given out when 2 kg of water cools from 60°C to 10°C. For water, C = 4200 J/ kg°C.

Solution

Heat lost or gained is given as $= mC\,\theta$

$$= 2 \times 4200 \times 50 = 420000 \text{ J.}$$

The instrument used for measuring the specific heat capacity is called the **calorimeter**. You may like to find out, what is a calorimeter, how it used?

Calorimeter

It is a cylindrical vessel of metal made usually of copper of small thickness. It is polished inside and outside to avoid heat loss due to radiation. It is kept inside an insulator. It has a lid of an insulating material like wood. This lid has two holes, one for the thermometer and the other for a stirrer. It is commonly used for determination of specific heats.

Calorimetry

The measurement of the quantity of heat (or the specific heat of a substance) is called calorimetry. This is because the unit of heat chosen is calorie. The basic principle of calorimetry is the conservation of heat, that is,

Heat lost by a hot body = Heat gained by a cold body in contact.

We now describe a method for measuring the specific heat of a solid.

Specific Heat of a Solid by the Method of Mixtures

The apparatus to measure specific heat of a solid was devised by Regnault. It consists of two parts, heater and the calorimeter (Figure 2.5). The heater consists of an air-chamber (A) surrounded by two coaxial cylinders (1).

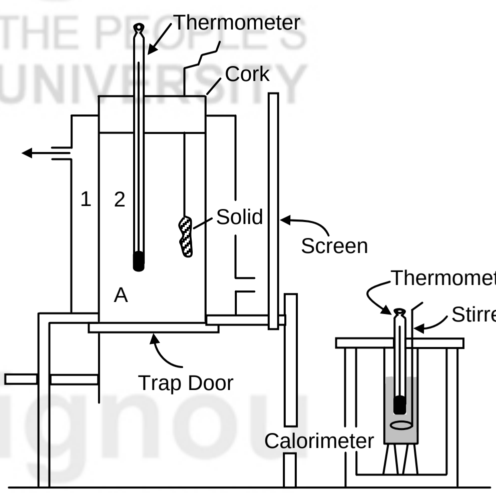

Figure 2.5 : Regnault Apparatus for the Determination of Specific Heat of Solids

In the space between the chamber and cylinder, steam is circulated continuously. The top of the air chamber is closed by a cork and the bottom by a trap door, through which the solid S is dropped in the calorimeter. A thermometer is also inserted through the cork to observe the temperature to which the solid is heated before being put in the calorimeter.

The calorimeter is a copper vessel polished on the outside. It is supported on an insulated stand so that heat losses due to radiation are minimized.

The heater is separated from the calorimeter by a wooden screen so that there is no direct exchange of heat between the heater and the calorimeter.

The following procedure is adopted :

The given solid in the form of a small piece is weighed and suspended in the heater. Steam is passed in the heater so that the solid is heated without actual contact with steam. The temperature of the heated solid is observed.

The empty dried calorimeter is weighed with the stirrer. Now some water is taken in the calorimeter and it is weighed again. The amount of water taken should be just enough to immerse the solid. The mass of water taken can be easily found.

The calorimeter is placed in its enclosure and the initial temperature of water is noted. When the solid has attained the steady maximum temperature of the heater, the calorimeter is pushed below the trap-door and the solid is quickly dropped into the calorimeter. The calorimeter is quickly moved back and the screen is lowered. The contents are well stirred and the final temperature is observed.

The following observations are recorded :

m = Mass of the solid

C = Specific heat capacity of the solid (to be determined)

θ = Temperature of the heated solid

m_1 = Mass of the calorimeter

C_1 = Specific heat capacity of the material of the calorimeter (given)

m_2 = Mass of water taken

C_2 = Specific heat capacity of water (given)

θ_1 = Initial temperature of water + calorimeter

θ_2 = Final temperature of the mixture

Heat lost by the solid = $mC\,(\theta - \theta_2)$

(The temperature of the solid falls from θ to θ_2)

Heat gained = Heat gained by calorimeter + Heat gained by water

$$= m_1 C_1 (\theta_2 - \theta_1) + m_2 C_2 (\theta_2 - \theta_1)$$

$$= (m_1 C_1 + m_2 C_2)(\theta_2 - \theta_1)$$

Since, Heat lost = Heat gained

So, $$mC(\theta - \theta_2) = (m_1 C_1 + m_2 C_2)(\theta_2 - \theta_1) \qquad \ldots (2.6)$$

From this the specific heat capacity C of the given solid can be calculated. Let us solve a few problems using this relation.

Example 2.2

A copper calorimeter of mass 120 g has the specific heat capacity of 420 J/kg°C. What is its heat capacity? 80 g of water at 15°C are poured in the calorimeter and it is heated to 55°C. Calculate the total heat supplied. Specific heat capacity of water is 4200 J/kg°C.

Solution

Heat capacity = $mC = 0.12 \times 420 = 50.4$ J/°C

Heat gained by calorimeter = $mC\,\theta = 0.12 \times 420 \times (55 - 15) = 2016$ J

Heat gained by water = $mC\,\theta = 0.08 \times 4200 \times 40 = 13440$ J

Total heat supplied = 2016 + 13440 = 15456 J

Example 2.3

A piece of metal 400 g is heated to 100°C and then quickly transferred to a copper calorimeter of mass 200 g and of specific heat capacity 380 J/kg°C containing 250 g of water at 25°C. The final temperature is found to be 40°C. Find the specific heat capacity of the metal. Specific heat capacity for water is 4200 J/kg°C.

Solution

Total heat gained = Heat gained by the calorimeter + Heat gained by water

$$= 0.2 \times 380 \times (40 - 25) + 0.250 \times 4200 \times (40 - 25)$$

$$= 1140 + 15750 = 16890 \text{ J.}$$

Heat lost by the metal = $mC\,\theta$

$$= 0.4 \times C \times (100 - 40) = 24\,C$$

$\therefore$ Heat lost = Heat gained

$\therefore$ $24\,C = 16890$

$\therefore$ $C = 703.75$ J/kg°C.

Example 2.4

A piece of metal of mass 400 g and of specific heat 126 J/kg°C heated to 150°C is dropped in 80 g of water at 20°C in a calorimeter of 150 g and of specific heat capacity 400 J/kg°C. Find the final temperature of the mixture. Specific heat for water is 4200 J/kg°C.

Solution

Let θ be the final temperature. Therefore, fall in temperature for the metal is (150 – θ) and the rise in temperature for water and calorimeter is (θ – 20).

$\therefore$ Heat lost by the metal = Heat gained by water + Heat gained by the calorimeter.

$$\therefore \quad 0.4 \times 126 \times (150 - \theta) = 0.08 \times 4200 \times (\theta - 20) + 0.150 \times 400\,(\theta - 20)$$

$$\therefore \quad 7560 - 50.4\,\theta = 336\,\theta - 6720 + 60\,\theta - 1200$$

$$\therefore \quad 7560 + 6720 + 1200 = 336\,\theta + 60\,\theta + 50.4\,\theta$$

$$\therefore \quad 15480 = 446.4\,\theta$$

$$\therefore \quad \theta = 34.7°\text{C.}$$

Note : The same method can be used to find the specific heat capacity C_2 of a liquid. In this case a solid of known specific heat capacity (C) is taken and instead of water, the given liquid is used. Then its specific heat capacity (C_2) can be found by using the relation.

$$mC\,(\theta - \theta_2) = (m_1 C_1 + m_2 C_2)\,(\theta_2 - \theta_1)$$

The specific heat of water has a high value, which can be used for many purposes.

Consequences and Advantages of High Specific Heat of Water

Due to the high specific heat of water, it can be used for heating as well as cooling purposes. When a given mass of water is heated through a given temperature, then because of high specific heat of water, it absorbs a large quantity of heat ($Q = mC\,\theta$). Because hot water has a large amount of heat, it can be used for fermentation, as it is capable of giving out large amount of heat. Water is also used in the radiators of cars for cooling purposes. When water is circulated around the heated engine, it absorbs heat. Due to its high specific heat it can absorb more heat than any other liquid for the same rise of temperature. For the same reason, wine and juice bottles are kept under water in cold countries to prevent them from freezing.

The climate in coastal areas influenced by the large body of water near them.

SAQ 2

(a) Calculate the amount of heat required to raise the temperature of 50 g of copper from 10°C to 60°C. (The specific heat capacity of copper is 0.39 J/g °C or 0.09 cal/g °C).

(b) How much heat is gained by 60 g of mercury when its temperature rises from 27°C to 60°C (Specific heat capacity of mercury is 1.39 J/g °C or 0.33 cal/g °C).

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

(d) A 200 g mass of a certain metal at 83°C is immersed in 300 g of water at 23°C. The final temperature is 33°C. Calculate the specific heat capacity of the metal.

2.4 TRANSMISSION OF HEAT

In general heat travels from one point to another *whenever there is a difference of temperatures.* Heat flows from a body at higher temperature to the body at lower temperature. Heat is transferred or propagated by three distinct processes : **conduction**, **convection** and **radiation**. *The net transfer of heat may take place by one or more of these distinct processes.*

2.4.1 Conduction

It is a process of transmission of heat from particle to particle, in the direction of the decrease of temperature without any *visible* movement of the particles themselves. This can be understood from the following simple experiment :

Hold one end of an iron rod and heat the other end. In the beginning the end which is held in the hand is not warm as it is at the room temperature, but soon this end becomes warmer and warmer till we are unable to hold the rod.

It is quite obvious that some heat has been transmitted (*conducted*) *along the rod from the heated end but the particles have not moved from their initial position*.

We now present a simple explanation of conduction.

When a rod is heated at one end, the molecules or atoms at the heated end receive more energy. Therefore, they start vibrating with increasing speed and amplitude. They strike the neighbouring particles more vigorously. The neighbouring particles also start vibrating with greater amplitude. This continues and the heat energy is transferred from particle to particle till this energy reaches the other end, which now becomes hot. Thus, conduction is a process in which the heat energy is transferred from particle to particle, without the particles leaving the mean positions but vibrating more vigorously with amplitudes which depend on the temperature. Thus, for conduction the medium is actively involved.

All substances conduct heat to some extent, but the rate at which heat is conducted is different for different substances. *The substances through which heat is easily conducted or for which the rate of conduction is large are* ***good conductors of heat****.*

The explanation of the good conductivity of the substances is based on the idea of free electrons, which are also known as 'conduction electrons'. The free electrons are present in the outermost orbits of atoms of the metals. When a metal is heated, the free electrons which are free to move in the body, begin to move faster and drift towards the cooler parts of the body and transfer the energy to the atoms on the way by collisions. Note that the electrons in the inner orbits are bound in the atom, therefore, they do not conduct heat.

The bad conductors do not have free electrons, therefore, they cannot conduct heat. Whatever little heat they can conduct is by vibrations of the molecules.

Comparison of Thermal Conductivities of Solids (Ingen-Hauz's Experiment)

A number of rods of different materials (Figure 2.6) but exactly of the same size (length and diameter) are inserted in the holes in the side of a vessel (trough). The portion of the rods projecting outside the vessel are dipped into molten wax. The rods are withdrawn and allowed to cool so that a uniform thickness of wax is formed.

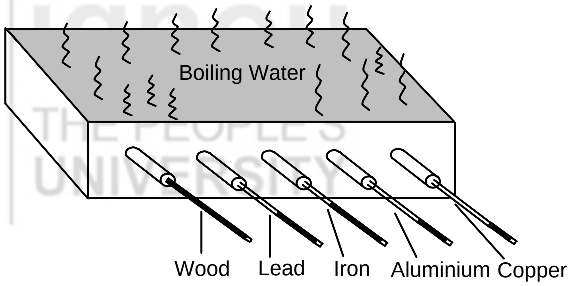

Figure 2.6 : Comparison of Thermal Conductivities

The rods are reinserted and boiling water is poured in the trough. The ends of the rod in the vessel are at the same temperature. As the heat is conducted along the rod, the wax begins to melt.

After a few minutes, 'steady state' is reached, that is no further melting of wax takes place. It is observed that for different rods, wax has melted up to different lengths, which shows that the conductivity of different substances is different. *That substance has the highest conductivity for which the wax melts to the maximum - length.*

2.4.2 Convection

It is a process by which liquids and gases are heated. It is a process of transmission of heat through the substance by the heated particles moving bodily from place to place. Since the particles in a solid are not free to move, therefore, transmission of heat by convection in solids is not possible.

You can do the following activity to observe convection in liquids.

Convection in Liquids

Take water in a large spherical flask and drop a single large crystal of potassium permanganate to the bottom. Take care to slide the crystal through a glass tube so that it does *not* colour the water on the way. Heat the flask slowly by using a narrow flame. Notice that an upward current of coloured water ascends from the place where heat is applied. This coloured stream reaches the *top* and spreads out. After a short time it circulates *down* the sides of the flask. This shows that convection current has been set up.

How do we explain occurrence of convection currents?

Explanation of Convection or Convection Currents

The layer of the liquid near the bottom of the vessel expands on heating. Therefore, its density decreases (as its mass remains constant). By Archimedes' principle it begins to rise upwards. Thus, a warm convection current moves upwards. Its place is taken up by the denser layer from the sides. This continues till the entire liquid is heated to the same temperature. The convection currents then stop and the density of the liquid is the same everywhere. (It should be noted that the liquids are always heated from below. If a liquid is heated near the *top*, then the upper layer which becomes lighter remains at the *top* and the entire liquid *cannot* be heated.)

The hot and cold regions on the earth's surface give rise to ocean currents. These are actually convection currents set up due to the difference of temperatures. According to the place of the origin these are hot or cold currents. In the equatorial region, water in the ocean is heated by the sun and it's density decreases. But, water in the ocean in the polar region remains cold and has higher density. Consequently currents of warm water flow towards the poles. These are 'hot currents'.

Below the surface of the ocean, currents of cold water flow from the poles to the equator. These are 'cold currents'. These currents control the temperature of the oceans.

2.4.3 Radiation

It is a matter of common experience that when we sit in the sun, we feel warm. When we sit *near* a fire we feel warm or if we hold our hand *below* a lighted bulb our hand feels warm.

How do we receive heat in the above cases? The process responsible for transmission of heat cannot be conduction, because the intervening medium (air) is a very poor conductor of heat and the heat from the sun travels *almost* the entire distance in vacuum. The process cannot be convection *because the hot air will always rise upwards and not sideways or downwards*.

Thus, there must be a third process by which heat can be transmitted *without the intervening medium* (which is not heated if it is present). This process of transmission of heat is known as **radiation**.

Transmission of energy by radiation takes place in the form of electromagnetic waves. These waves can travel in vacuum and travel with the speed of light ($3 \times 10^8\ \mathrm{ms}^{-1}$). This heat is known as the radiant heat or thermal radiation. It is also known as infra-red radiation because 'heat waves' occupy a place till the electromagnetic spectrum just above the red colour in the visible spectrum .

Detection of Radiant Heat

A Blackened-bulb Thermometer

A thermometer of large bulb is arranged at a certain distance from the source of heat (a candle or a bulb). Its temperature is noted after a certain time. Now blacken the bulb of the thermometer with soot blacken and observe its temperature after the same time. Now the thermometer shows a greater rise in temperature. Therefore, a blackened bulb thermometer can be used to detect "radiation" because the blackened-bulb is able to absorb the entire incident heat.

A good absorber of heat is also a good radiator of heat, that is, a body which can absorb a large amount of heat, when heated to a high temperature would also radiate large amount of heat in a given time. The absorbing power depends on the nature of the surface and the area on which heat is incident. Similarly. When the body is heated to a high temperature, it would radiate heat which would depend on the temperature, nature of the surface and the area of the surface.

A body which is capable of absorbing the entire incident heat is a perfect black body. Such a body does not reflect or transmit heat. In practice a body covered with lamp black (soot) is taken as a black body. It can absorb up to 95% of the incident heat.

A black body is not only a good absorber of heat but also a good emitter. Similarly, a polished surface is not only a poor absorber (as it reflects most of the heat) but also a poor radiator of heat.

Take two identical thermometers and blacken the bulb of one of the thermometers. Arrange them parallel to each other and place a source of heat (a bulb or a candle) midway between their bulbs. It will be found that there is a rise in temperature of both the thermometers but the thermometer with blackened bulb not only shows greater rise in the temperature but its rate of rise in temperature is also rapid. This shows that a black body is a good absorber of heat.

Practical Uses of Radiation

(i) In summer, we wear light coloured or white clothes because they absorb very little heat and reflect a very large amount of heat so that we feel quite comfortable.

(ii) Hot water pipes, radiators of cars are painted black so that they radiate heat quickly and the temperature is kept within limits.

(iii) Electric radiators use highly polished reflectors so that a large amount of heat produced is reflected in the room.

SAQ 3

(a) What is conduction? Explain with the help of an example the process of heat transmission by conduction.

(b) Describe an experiment to compare the conductivities of different solids.

(c) What is convection? Describe an experiment for its illustration.

(d) What is radiation? How is it possible to detect radiant heat?

2.5 CHANGE OF STATE

You know that one and the same substance can exist in three different states – solid, liquid and gas. For example, ice is the solid state, water is the liquid state and steam is the gaseous state of the same substance. We generally observe that when heat is supplied to a substance, its temperature increases and when heat is withdrawn *from* a substance, its temperature decreases. What do you observe when you heat ice or water? Ice melts into water and water turns into steam. We say that a change of state has occurred. We now discuss this phenomenon. The supply of heat to a substance produces a change *from* the solid state to the liquid state *or from* the liquid state to the gaseous state.

Conversely the withdrawal of heat *from* a substance causes it to change from the gaseous state to the liquid state or from the liquid state to the solid state.

The change from the solid to the liquid state is known as **fusion** or **melting** or **liquefaction**. While the reverse change *from* the liquid to the solid state is called **solidification** or **freezing**. The change from liquid to vapour is known as **vaporization** and the reverse process as **condensation**.

The constant temperature at which a substance changes from solid to liquid state is called its ***melting point****.*

We now describe the method of determining the latent heat of fusion of ice.

Determination of the Latent Heat of Fusion of Ice by the Method of Mixture

A clean and dry calorimeter with a stirrer is weighed empty. Then water is filled in about two-thirds of the container. It is weighed again and its temperature is noted. To minimize the heat exchange between the calorimeter and the surroundings, the calorimeter is placed on an insulated support in an enclosure.

A few pieces of ice which have been dried by a blotting paper are dropped in the calorimeter. The mixture is stirred until the ice is completely melted and the final steady temperature is noted. The calorimeter and contents are weighed again and the following observations are taken.

m_1 = Mass of the calorimeter (empty) + stirrer

m_2 = Mass of the calorimeter and stirrer + water

$\therefore$ m_w = Mass of water taken = $m_2 - m_1$

θ_1 = Initial temperature

θ_2 = Final temperature of the mixture

m_3 = Mass of the calorimeter and stirrer +water + ice

m = Mass of the ice added = $m_3 - m_2$

C_1 = Specific heat capacity of the calorimeter (given)

C = Specific heat capacity of water (given)

Heat gained by ice = Heat required to melt it + Heat required to raise the temperature of melted ice from 0^oC to θ_2

$$= mL + mC\,(\theta_2 - 0)$$

Heat lost = Heat lost by calorimeter and water in it.

$$(m_1 C_1 + m_w C)\,(\theta_1 - \theta_2)$$

Heat lost = Heat gained

$$L = \frac{(m_1 C_1 + m_w C)\,(\theta_1 - \theta_2) - mC\,\theta_2}{m}$$

From this the latent heat of fusion of ice can be calculated.

Example 2.5

10 g of ice at 0°C are added to 100 g of water at 20°C. If the final temperature of mixture is 10.9°C, find the latent heat of fusion of ice. Specific heat of water is 4200 J/kg°C.

Solution

Heat is gained by ice in two stages. To melt it into water at 0°C, the heat absorbed is mL. Then to raise its temperature to 10.9°C, the heat absorbed is $mC\,(10.9 - 0)$.

$\therefore$ Total heat gained by ice = $mL + mC\,(10.9 - 0) = 0.01 \times L + 0.01 \times 4200 \times 10.9$

Heat lost by water = $0.1 \times 4200 \times (20 - 10.9) = 3822$ J

$\therefore$ Heat gained = Heat lost

$\therefore$ $0.01\,L + 0.01 \times 4200 \times 10.9 = 0.1 \times 4200 \times 9.1$

$\therefore$ $L = 336420$ J/kg.

Example 2.6

How much ice at 0°C should be added to 250 g of water so as to reduce its temperature from 35°C to 25°C. Latent heat of ice is 3360 J/g and specific heat of water is 4.2 J/g ^oC.

Solution

Let x g be the mass of ice. Heat is gained by ice in melting and then the temperature of melted ice is increased from 0°C to 25°C.

Heat gained by ice $x \times 3360 + x \times 4.2 \times 25 = 3465\, x$ J

Heat lost by water = $250 \times 4.2 \times 10 = 10500$ J

Hence, $3465\, x = 10500$

$$\therefore \qquad x = \frac{10500}{3465} = 3.03 \text{ g}$$

Example 2.7

(a) A beaker contains 100 g of water at 20°C. The beaker and its contents are heated till 50 g of water are boiled off. Calculate

(i) the heat required to raise the temperature of water to 100°C.

(ii) the total amount of heat used in the experiment. (Neglect the heat taken up by the beaker.) The specific heat capacity of water is 4.2 J/g °C. The specific latent heat of steam is 2260 J g^{-1}.

(b) The water in the beaker is cooled to 50°C and then 50 g of copper at 18°C are added. The contents are stirred till a final equilibrium temperature is reached. Calculate this final temperature. The specific heat capacity of copper is 0.39 J/g°C. Give your answer to the nearest half degree.

Solution

(a) (i) Heat required to raise the temp of water to 100°C

$$= mC\,\theta = 100 \times 4.2 \times (100 - 20) = 33600 \text{ J}$$

(ii) Heat required to evaporate 50 g of water

$$= mL = 50 \times 2260 = 113000 \text{ J}$$

Total heat = 33600 + 113000 = 146600 J

(b) Let θ be the final temperature

Heat lost by water = Heat gained by copper

$$50 \times 4.2\,(50 - \theta) = 50 \times 0.39 \times (\theta - 18)$$

$$10500 - 210\,\theta = 19.5\,\theta - 351$$

$$229.9\,\theta = 10851$$

$$\therefore \qquad \theta = 47.28°\text{C}.$$

Example 2.8

1 kg of molten lead at its melting point of 327°C is dropped into 1 kg of water at 20°C. Assuming no heat is lost, calculate the final temperature of water.
Take : Specific heat capacity of lead = 130 J/kg °C, Latent heat of fusion of lead = 27000 J/kg, Specific heat capacity of water = 4200 J/kg °C.

Solution

Let θ be the final temperature of water

Heat lost by lead

$$= mL + mC\,(327 - \theta)$$

$= 1 \times 27000 + 1 \times 130 \times (327 - \theta)$

$= 27000 + 42510 - 130\ \theta$

$= 69510 - 130\ \theta$

Heat gained by water

$= mC\ (\theta - 20)$

$= 1 \times 4200\ (\theta - 20)$

$= 4200\ \theta - 84000$

Since, Heat lost = Heat gained

$69510 - 130\ \theta = 4200\ \theta - 84000$

$4330\ \theta = 153510$

$\theta = 35.5°C$

Consequences and Advantages of High Latent Heat of Water

(a) The ponds and lakes in cold countries freeze at a very slow rate. Even after the temperature of water becomes 0°C, the freezing does not start immediately as for every one gram of water to be frozen 336 J will have to be removed. Because of this value of latent heat the freezing will take place very slowly and aquatic life can survive in it.

(b) We cool our drinks by adding pieces of ice. Since for melting one gram of ice, 336 J are removed from the drink, therefore, it loses considerable amount of heat and can be cooled by using a small quantity of ice.

(c) The water in rivers is due to snow which melts on mountaintops. Since the latent heat of ice is high, therefore, ice melts at a very slow rate and we can get water in rivers throughout the year. If the latent heat of ice were small, then the ice would melt quickly leading to floods in rivers.

Effect of Pressure on Melting Points of Solids

Let V_f and V_i be the final and initial volumes of a solid. In case of most solids, the final volume is more than the initial volume but there are certain exceptions like ice, bismuth, gallium, etc. whose volume decreases when melting takes place.

(a) When melting point is accompanied with an expansion, i.e. $V_f > V_i$, e.g. wax; dp/dT is positive. Hence the melting point increases with the increase in pressure increment or vice-versa.

(b) When melting point is accompanied with compression, i.e. $V_f < V_i$, e.g. ice; dp/dT is negative and hence increase in pressure leads to a decrease in melting point and vice-versa.

The above explanation is based on Clausius-Clapeyron equation.

If dp = change in pressure,

and dT = change in temperature,

L = latent heat,

V_f = final volume,

V_i = initial volume, and

T = temperature at which phase change takes place, then

$$\frac{dp}{dT} = \frac{L}{T\ (V_f - V_i)} \qquad \ldots (2.7)$$

Heat is absorbed by a substance during melting and an equal amount of heat is given up by the substance during solidification without showing any rise or fall of temperature.

Similarly, a liquid absorbs heat during vaporization and an equal amount of heat is given by it during condensation without showing any rise or fall of temperature.

The amount of heat required to change the state of a substance without any change in temperature is called the latent heat.

If this heat is expressed for a unit mass, then it is known as the **specific latent heat**.

Specific Latent Heat of Fusion

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

Its units are : joules per kg (J/kg),

Specific latent heat of ice (L) = 336000 J/kg = 336 kJ/kg.

Specific Latent Heat of Vaporization

The specific latent heat of vaporization of a substance is the quantity of heat required to convert unit mass of the substance from the liquid to the vapor state without change of temperature.

Specific latent heat of steam (L) = 2260000 J/kg (2260 kJ/kg)

Effect of Pressure on Boiling Point of Liquids

When a liquid is boiling it temperature T and pressure P change, i.e. change of state takes place and it gets converted into vapour. Thus, V_f the final volume, is always found to be much larger than V_i, the initial volume. Thus, dP/dT from Clausius-Clapeyron (Eq. (2.7)), is positive and hence the boiling point is found to be increasing with increasing pressure and vice-versa.

At high latitudes the atmospheric pressure is lower. Therefore, water boils at temperatures less than 100°C. By using pressure cooker, pressure is increased so that a boiling point of more than 100°C may be attained which accelerates the cooking process.

SAQ 4

(a) What mass of ice at 0°C will be required to cool 900 g of water from 35°C to 10°C. Assume all the ice used melts (The specific heat capacity of water is 4.2 J/g °C, the specific latent heat of fusion of ice is 336 J/g.)

(b) What mass of ice at 0°C needs to be added to 100 g of water at 17°C to lower the temperature to 0°C? (Specific latent heat of ice is 340 J/g, specific heat capacity of water is 4.2 J /g °C.)

(c) Calculate the total amount of heat required to convert 100 g of ice at – 10°C completely into steam at 100°C. (Specific heat capacity of ice = 2100 J/kg °C; specific heat capacity of water = 4200 J/kg °C; specific latent heat of fusion of ice = 336,000 J/kg; specific latent heat of vaporization of water = 2260,000 J/kg.)

2.6 SUMMARY

Let us summarise what you have learnt in this unit.

In this unit, we have learnt the fundamentals of thermal expansion of liquid. We all know that most materials expand when their temperatures increase. Rising temperatures make the liquid expand. A completely filled and tightly capped bottle of water cracks when it is heated, but you can loosen a metal bar lid by running hot water over it. These are the few examples of thermal expansion. A detailed discussion has also been carried out for the anomalous (unusual) expansion of water. You have also been introduced to the measurement of heat.

The three modes of heat transfer are conduction, convection and radiation. Conduction occurs within a body or between two bodies in contact. Convection depends on motion of mass from one region of space to another. Radiation heat transfers by electromagnetic radiation, such as sunshine with no need for matter to be present in the space between bodies. You have also understood the concept of heat conduction that comes into play whenever two objects at different temperatures are placed in contact, whereas convection is the transfer of heat by mass motion of a fluid from one region of space to another. Radiation is the transfer of heat by electromagnetic waves such as visible light, infrared and ultraviolet radiation.

The compound H_2O exists in the solid phase as ice, in the liquid as water, and in the gaseous phase as steam. A transmission from one phase to another is called change of state. You have also learnt the basic concepts of change of state.

2.7 KEY WORDS

Calorie : It is defined as the amount of heat required to raise the temperature of one gram of water from 14.5°C to 15.5°C.

Coefficient of Apparent Expansion : It is defined as the apparent increase in volume per unit original volume for 1°C rise in temperature.

Coefficient of Real Expansion : It is defined as the real increase in volume per unit original volume from 1°C rise in temperature.

Conduction : Conduction is a process of heat transfer by which heat flows from a region of higher temperature to a region of lower temperature within a medium or between different media in direct physical contact by molecular vibration.

Convection : Convection is a process by which heat is transmitted through a substance from one point to another point due to the bodily motion of the heated particles at the substance.

Radiation : The flow of energy by virtue of electromagnetic wave even through vacuum is called radiation. In nature maximum heat transfer takes place from radiation.

Heat : It is a form of energy, which produces in us the sensation of warmth.

Temperature : The degree of hotness of a body is called its temperature.

Specific Heat : It is defined as the amount of heat required to raise the temperature of the unit mass of the substance through 1°C (or 1 K).

Latent Heat : The latent heat of a substance may be defined as the quantity of heat required to change the unit mass of the substance completely from its one state to another at constant temperature. The unit of latent heat is J kg^{-1} in SI units.

Latent Heat of Fusion : The latent heat of fusion of solid may be defined as the amount of heat required to change the unit mass of the solid to its liquid state at its melting point.

Latent Heat of Vaporisation : The latent heat of vaporisation of a liquid may be defined as the amount of heat required to change the unit mass of the liquid to its vapour state at its boiling point.

2.8 ANSWERS TO SAQs

SAQ 1

(a) $$\Delta V_{\text{glass}} = \gamma_{\text{glass}} \times V_0 \times \Delta T$$

$$= 1.2 \times 10^{-5} \times 200 \times (100 - 20)$$

$$= 0.19 \text{ cm}^3$$

The increase in volume of the mercury is

$$\Delta V_{\text{mercury}} = \gamma_{\text{mercury}} \times V_0 \times \Delta T$$

$$= 18 \times 10^{-5} \times 200 \times (100 - 20)$$

$$= 2.9 \text{ cm}^3$$

The volume of mercury that overflows is

$$\Delta V_{\text{mercury}} - \Delta V_{\text{glass}} = 2.9 - 0.19$$

$$= 2.7 \text{ cm}^3$$

(b) Here the volume of air in the flash remains the same. This is only possible when the expansion of flask is exactly the same as the expansion of mercury in the flask.

Let x be the volume of mercury in the flask.

Here volume of flask = V = 1 litre = 1000 cm^3

Expansion of flask = Expansion of mercury

$$V \gamma_g T = x \gamma_m T$$

$$\therefore \quad x = V \frac{\gamma_g}{\gamma_m}$$

$$= 1000 \times \frac{27 \times 10^{-6}}{1.8 \times 10^{-4}}$$

$$= 150 \text{ cm}^3$$

(c) Here $\gamma = 49 \times 10^{-5}\ {}^{\circ}\text{C}^{-1}$; $\Delta T = 30^{\circ}\text{C}$

Let there be m grams of glycerine and its initial volume be V. Suppose the volume of glycerine becomes V' after a rise of temperature of 30°, then

$$V' = V\,(1 + \gamma\,\Delta T)$$

or, $V' = V\ (1 + 49 \times 10^{-5} \times 30)$

$= 1.0147\ V$

Initial density of glycerine $= \rho = \dfrac{m}{V}$

Final density of glycerine $= \rho' = \dfrac{m}{V'}$

$$= \frac{m}{1.047\ V}$$

$$= \frac{\rho}{1.047} = 0.9855\ \rho$$

Therefore, fractional change in the volume of density of glycerine

$$= \frac{\rho - \rho'}{\rho}$$

$$= \frac{\rho - 0.9855\ \rho}{\rho}$$

$$= 0.0149$$

(d) Initial density of mercury at ^{o}C $= \rho = \dfrac{m}{V}$

Final density of mercury at 50^{o}C $= \rho' = \dfrac{m}{V'}$

Now $\dfrac{\rho}{\rho'} = \dfrac{V'}{V}$

We know $V' = V\ (1 + \gamma\ \Delta T)$

Now $\dfrac{\rho}{\rho'} = \dfrac{V\ (1 + \gamma\ \Delta T)}{V}$

So, $\rho' = \dfrac{\rho}{1 + \gamma\ \Delta T}$

$= \rho\ (1 - \gamma\ \Delta T)$ (Since, γ is very small)

$= 13.6\ (1 - 1.82 \times 10^{-4} \times 50)$

$= 13.48$ g cm^{-3}

SAQ 2

(a) Here $m = 50$ g, specific heat (s) = 0.39 J/g^{o}C, $\Delta T = (60 - 10) = 50^{o}$C

We know $H = m\ s\ \Delta T$

$= 50 \times 0.39 \times (60 - 10) = 975$ J

(b) Here $m = 60$ g, specific heat $s = 1.39$ J/g^{o}C

$\Delta T = (60 - 27) = 33^{o}$C

We know $H = m\ s\ \Delta T$

$= 60 \times 1.39 \times 33 = 2752.2$ J

(c) Let the final temperature be t.

We know that, Heat loss = Heat gain

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

$$\Rightarrow \quad 4t - 100 = 80 - t$$

$$\Rightarrow \quad 5t = 180$$

$$\Rightarrow \quad t = 36^{o}\text{C}$$

So, final temperature is 36^{o}C

(d) Here m = mass of metal = 200 g = 0.2 kg

$t_1 = 83^{o}\text{C}$

M = mass of water = 300 g = 0.3 kg

$t_2 = 23^{o}\text{C}$

Final temperature $t = 33^{o}\text{C}$

Let C_1 be the specific heat of the metal.

Specific heat of water C_2 = 4200 J/kg ^{o}C

We know, Heat loss by metal = Heat gain by water

$$m\, C_1\, (t_1 - t) = M\, C_2\, (t - t_2)$$

$$\Rightarrow \quad 0.2 \times C_1\, (83 - 33) = 0.30 \times 4200 \times (33 - 23)$$

$$\Rightarrow \quad C_1 = 1260 \text{ J/kg}\,^{o}\text{C}$$

SAQ 3

Refer to preceding text.

SAQ 4

(a) Let the required mass of ice be m gram;

M = mass of water = 900 g

L = specific latent heat of fusion of ice = 336 J/g

C_1 = specific heat capacity of water = 4.2 J/g ^{o}C

Heat gained by ice = Heat lost by water

$$m\,L + m \times C_1 \times (10 - 0) = M \times C_1 \times (35 - 10)$$

$$\Rightarrow \quad m \times 336 + m \times 4.2 \times 10 = 900 \times 4.2 \times 25$$

$$\Rightarrow \quad m = 250 \text{ g}$$

So, 250 g ice will be required.

(b) Let the required mass of ice be m gram

M = mass of water

L = specific latent heat of ice = 340 J/g

C_1 = specific heat capacity of water = 4.2 J/g ^{o}C

Heat gained by ice = Heat lost by water

Temperature of water = t^{o}C

$$m\,L = M\;C_1\;(t - 0)$$

$$m \times 340 = 100 \times 4.2 \times (17 - 0)$$

$\Rightarrow \qquad m = 21$ g

So, 21 g of ice will be required.

(c) Required heat = Heat required to lower the temperature of ice from – 10°C to 0°C + Melt the ice at 0°C to water at 0°C + increase the temperature of water from 0°C to 100°C + convert water 100°C to steam at 100°C.

Mass of ice = 100 g = 0.1 kg

Required heat

$$= 0.1 \times 2100 \times [0 - (-10)] + 0.1 \times 336000 + 0.1 \times 4200 \times (100 - 0)$$
$$+ 0.1 \times 2260000$$
$$= 303700 \text{ J} \quad \text{or} \quad 303.7 \text{ kJ}$$