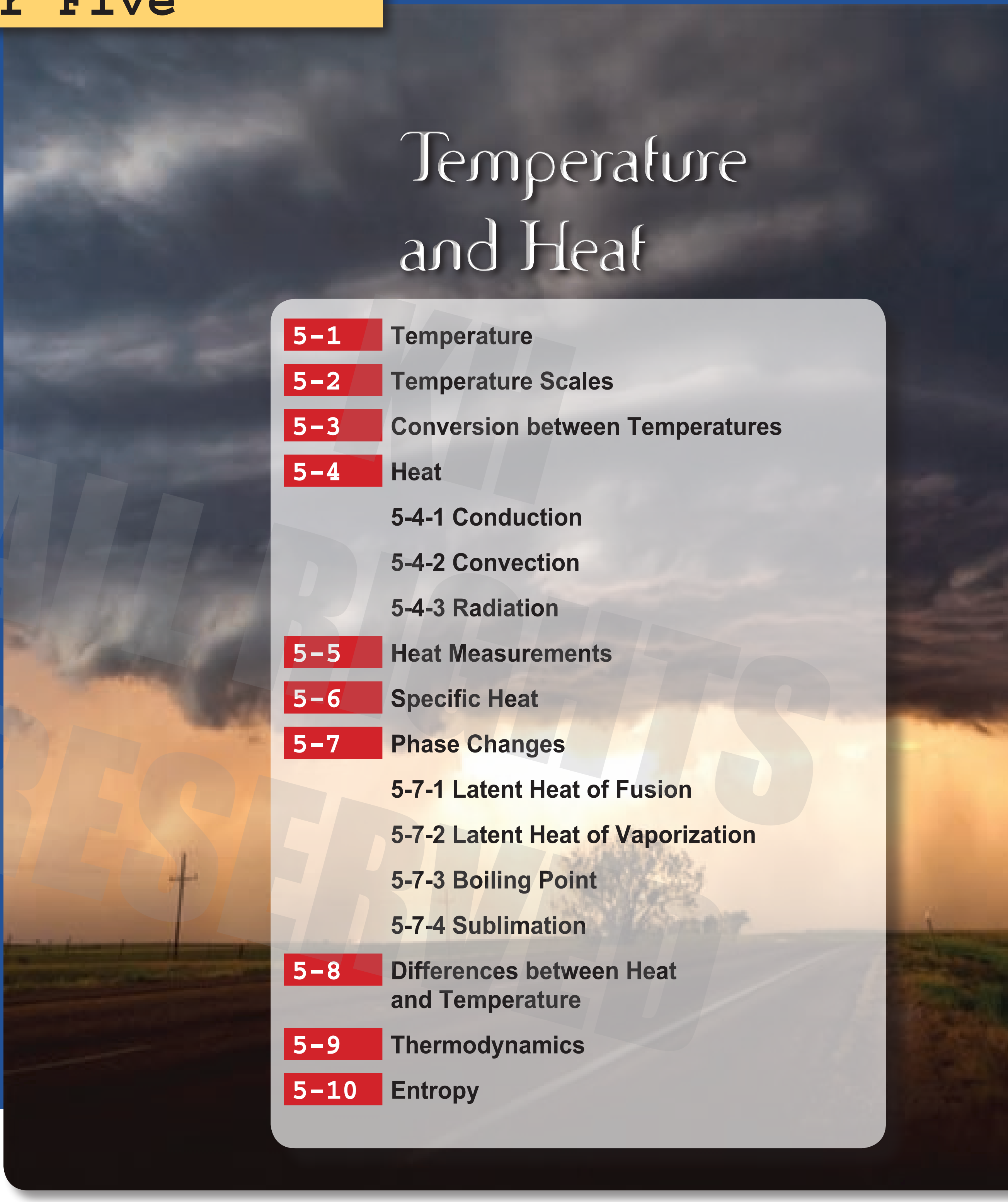
Chapter Five



Temperature and Heat

TEMPERATURE and heat are common terms that we use in our everyday lives. Some people think that temperature and heat are the same, but they are not. They are rather related to each other but not the same. It is important to know the difference between temperature and heat. In order to fully discuss and understand the difference between temperature and heat, we will first study temperature and heat.

5-1 Temperature

Temperature is not energy, but a measure of the average kinetic energy of the disordered motion of atoms and molecules of a substance. Thermometer is an instrument that is used to measure temperature in different units because it contains a **thermometric substance**, meaning a substance, like mercury, whose physical property changes with change in temperature. If the temperature is measured in degrees Kelvin, then the temperature value is directly proportional to the average kinetic energy of the molecules in the substance being measured. It means that if the Kelvin temperature of the substance is doubled, then the average kinetic energy of the molecules of the substance is also doubled.

If the average kinetic energy of the molecules of the substance goes up, meaning rise in temperature, the average speed of the molecules increases. Also, if the average kinetic energy of the molecules of the substance decreases, meaning drop in temperature, the average speed of the molecules of the substance goes down.

In fact, temperature is a property of a system, which determines whether or not heat is transferred to or from an object. In a qualitative manner, temperature can be described as the determination of whether a substance is warm or cold. If a thermometer is placed in thermal contact with a substance, and when thermal equilibrium is reached, a quantitative measure of the temperature of the substance can be determined.

5-2 Temperature Scales

A temperature scale is used for measuring temperature and the construction requires two reference temperatures. The reference temperatures are referred to as the ***fixed temperatures*** or ***fixed points***. There are upper fixed and lower fixed points. The upper fixed point is the steam point and is the temperature at which steam forms from pure water, boiling at normal atmospheric pressure. The lower fixed point is the temperature of a mixture of ice and water at normal atmospheric pressure.

The Fahrenheit and Celsius (centigrade) are more commonly used temperature scales. The Celsius scale was named after the inventor Anders C. Celsius in 1742. The Celsius scale has the upper fixed or steam point of 100° (100 degrees) and the low fixed or ice point of 0° (0 degrees), which means that the boiling point of water in the Celsius scale is 100°C and the freezing point of water is 0°C. There are 100 equal divisions between 0° and 100°, such that each division is one degree Celsius (1°C). The interval between 0°C and 100°C is called the ***fundamental interval***. Fundamental interval is a range of numbers between lower fixed point and upper fixed point. Figure 5.1 shows mercury in a glass thermometer.

The Fahrenheit scale was developed about 1715 by a German named Gabriel D. Fahrenheit. The development of the Fahrenheit scale was based on the upper fixed or steam point of 212° and the lower fixed or ice point of 32°. The interval between 32° and 212° is divided into 180 equal divisions. Each division is 1°F. In the Fahrenheit scale, the boiling point of water is 212°F and the freeing point of water is 32°F.

The Kelvin scale was developed by William Thompson, Lord Kelvin (1824–1907). The **Kelvin scale**, also called the ***absolute*** *scale* is abbreviated *K* (metric system). This scale was developed

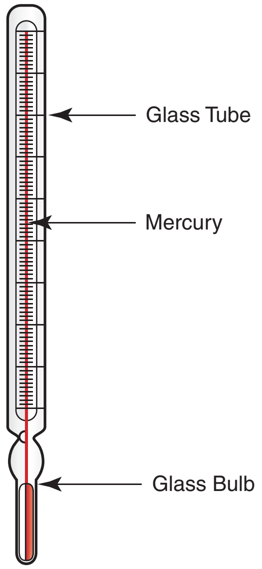


Figure 5.1 A mercury-in-glass thermometer. As temperature changes, the mercury in the glass bulb contracts or expands and the mercury level in the glass tube falls or rises.

when it was discovered that the temperature below which nothing can be cooled is 273°C. This reminds us that the temperature is proportional to the average kinetic energy of molecules. Cooling substances gradually immobilize molecules and reduce their kinetic energy until they are theoretically depleted of kinetic energy. This occurs at −273°, also called *absolute zero*. Absolute zero temperature can be determined when a graph of gas volume versus temperature is plotted (figure 5.2). In such an experiment to determine the absolute zero, the volumes or length of gas column is read at various temperatures. A straight-line graph is obtained, which shows that the rise in temperature is proportional to the increase in volume. By extending the graph until it meets the temperature axis, the point where the extended time meets the temperature axis, (−273°C) is the temperature where the volume would theoretically become zero. This is the absolute zero temperature.

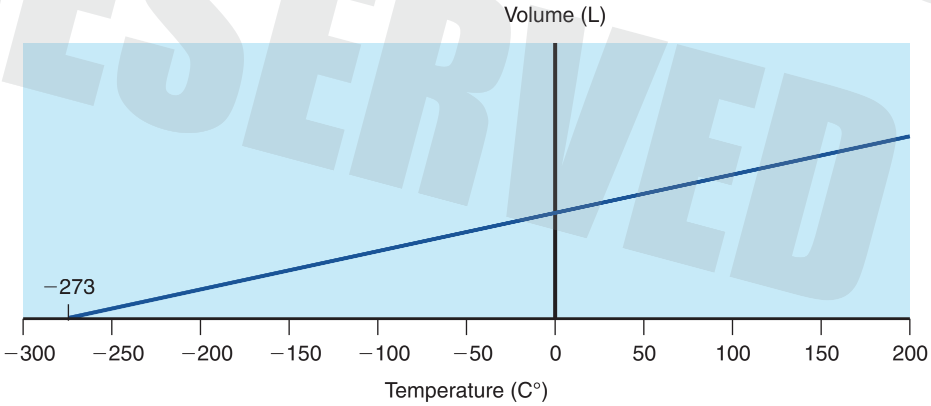


Figure 5.2 Graph of volume versus temperature at constant pressure.

In the Kelvin scale, temperatures are measured in degrees, the same size as degrees Celsius. The scale does not have negative values. Remember that no degree sign (°) is attached to Kelvin temperatures. 0 K is the same as −273°C. The freezing point (ice point) of water in the Kelvin scale is 273 K and 373 K is the boiling point (steam point) of water. The Celsius and the Kelvin scales are related as shown below.

$$\mathbf{T_K = T_C + 273.}$$

Figure 5.3 shows the three temperature scales.

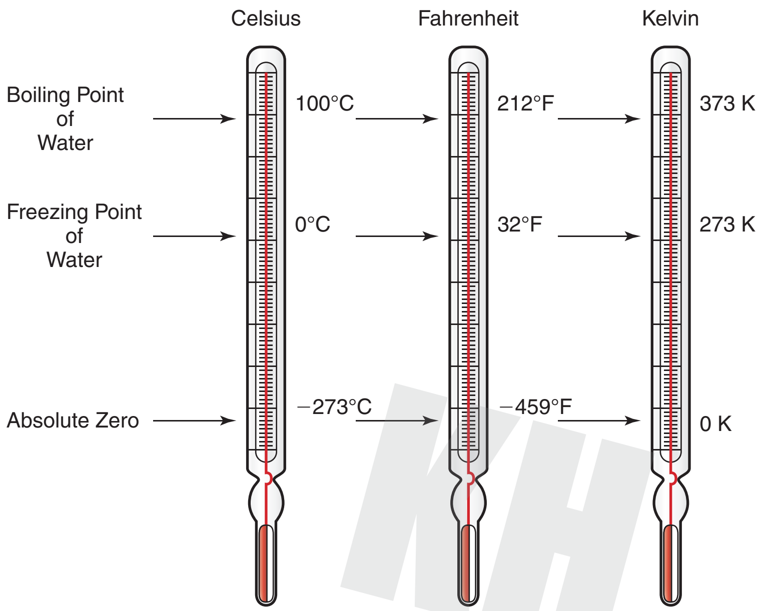


Figure 5.3 The Celsius, Fahrenheit and Kelvin temperature scales.

5-3 Conversion between Temperatures

Fahrenheit temperatures can be converted to Celsius temperatures and vice versa. The Celsius scale has 100° between the lower fixed point and the upper fixed point and the Fahrenheit scale has 180° between the lower and upper fixed points. Each Celsius degree is 180/100 or 9/5 of each Fahrenheit degree and each Fahrenheit degree is 100/180 or 5/9 of each Celsius degree. The equations for converting between Celsius and Fahrenheit are

$$°F \rightarrow °C \quad T_C = 5/9\,(T_F - 32°) \qquad \text{equation 5.1}$$

$$°C \rightarrow °F \quad T_F = (9/5\,T_C) + 32° \qquad \text{equation 5.2}$$

To convert Celsius scale to Kelvin scale, the following equation is used:

$$°C \rightarrow K \quad T_K = T_C + 273\text{ K}. \qquad \text{equation 5.3}$$

So, 40°C is the same as

$$T_K = 40°C + 273$$
$$= 313\text{ K}.$$

$$K \rightarrow °C \quad T_C = T_K - 273°C$$

EXAMPLE 5.1

QUESTION:

The air temperature outside is 60°C. What are the equivalents in (a) Fahrenheit and (b) Kelvin scales?

SOLUTION:

a. The Fahrenheit scale equation is

$$T_F = 9/5\,T_C + 32$$
$$= (9/5\;60) + 32$$
$$= 140°F.$$

b. The Kelvin scale equation is

$$\begin{aligned} T_K &= T_C + 273 \\ &= 60 + 273 \\ &= 333 \text{ K.} \end{aligned}$$

EXAMPLE 5.2

QUESTION:

The temperature in your laboratory is 75°F. What are the equivalents in (a) Celsius scale and (b) Kelvin scale

Answer: a. 23.9°C.

b. 296.9 K.

5-4 Heat

Heat, unlike temperature, is the measure of the energy in a substance. Heat is measured in joules or other energy units. Adding heat to a substance means adding energy to a substance. The added heat (energy) increases the kinetic energies of the atoms and molecules of the substance. If the heat (energy) is used to change the state of the substance, say by melting it, then the added energy is used to break the bonds between the molecules rather than changing their kinetic energy.

If heat (energy) is added to a substance, either the temperature of the substance increases or a change of state of the substance occurs. For the temperature of the substance to rise, the added heat (energy) will make the average kinetic energy of the atoms and molecules of the substance to increase as the atoms and molecules of the substance move faster than before the heat was added. This increase in average kinetic energy is registered as a number called temperature that changes proportionally with it.

If the substance changes its state (phase), for example, if the substance is ice, it will melt into water. In this case, there is no rise in temperature, because the temperature of the ice at the exact time before melting and the temperature of the water at the exact time after melting are the same. Heat is absorbed, but it is not **used to alter the temperature, rather** used to change the bonding between the molecules. If the water is to freeze back to water, heat (energy) is released.

QUESTION:

Why is there a heat (energy) change during phase change without change in temperature?

ANSWER:

During a phase change, heat (energy) is used to alter the bonding between the molecules of the substance. In the case of melting, added energy is used to break the bonds between the molecules. In the case of freezing, energy is subtracted as the molecules bond to one another. These energy exchanges are not changes in kinetic energy. They are changes in bonding energy between the molecules. The temperature remains the same, as shown in Table 5.1.

TABLE 5.1 PHASE CHANGES SHOWING NO CHANGES IN TEMPERATURE

Change of phase	Meaning of phase change	Heat motions during phase change	Change in temperature during phase change
Solid to liquid	Melting	As the solid melts, heat moves into the solid.	No temperature change
Liquid to solid	Freezing	As the liquid freezes, heat moves out of the liquid.	No temperature change
Liquid to gas	Boiling, vaporization, or evaporation	As the liquid boils, vaporizes, or evaporates, heat moves into the liquid.	No temperature change
Gas to liquid	Condensation	As the gas condenses, heat moves out of the gas.	No temperature change
Solid to gas	Sublimation	As the solid sublimates, heat moves into the solid.	No temperature change

Heat is transferred from an area of higher temperature to area of low temperature through **conduction**, **convection,** and **radiation**.

5-4-1 Conduction

Conduction means that heat flows through materials. When a substance is hot, the molecules move at a high speed and have more energy than the molecules of a cold substance. For example, a cold teaspoon placed in a hot teacup soon becomes equally hot. This is due to heat conduction. The heat conduction from an area of high temperature to an area of low temperature occurs when the molecules in motion in the high temperature area interact or collide with the slow moving molecules in the area of low temperature. Now, the molecules in the low temperature region acquire high temperature and the process continues. The thermometers show decrease of temperature from hotter to colder area (Figure 5.4).

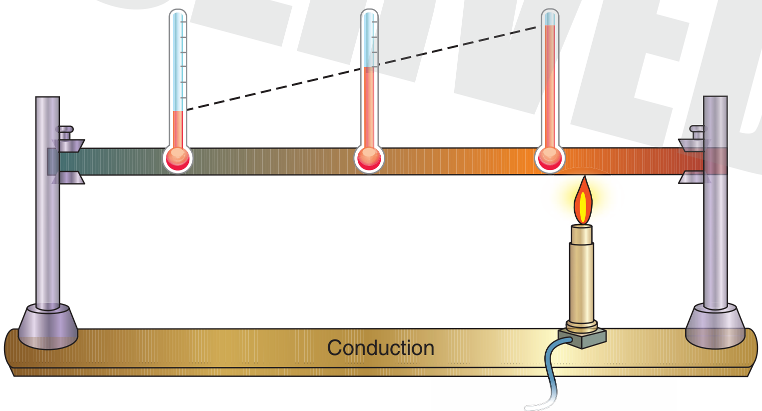


Figure 5.4 Heat Conduction.

Some substances such as metals allow heat to pass through them and are therefore called ***heat conductors***, while other substances, such as wood, glass, and rubber do not allow heat to pass through them. They are called ***poor heat conductors* or *insulators***. You may have observed that most cookware is made from copper because it is a very good conductor of heat.

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Cookware handles are made of wood or plastics because they are insulators. Some insulators such as Styrofoam have pockets of air spaces that prevent heat from passing through.

5-4-2 Convection

The **convection** of heat occurs in liquids and gases as molecules travel in currents. The convection currents are triggered when cold water is heated. The heating sets the molecules in motion, making the water warm and less dense. As the water becomes less dense, it rises and the heavier, or denser, cold water sinks to replace the rising, less dense water. In this way, a convection current is produced and heat is transferred (figures 5.5 and 5.6). Heat transfer by convection current is a lot easier in gases than in liquids

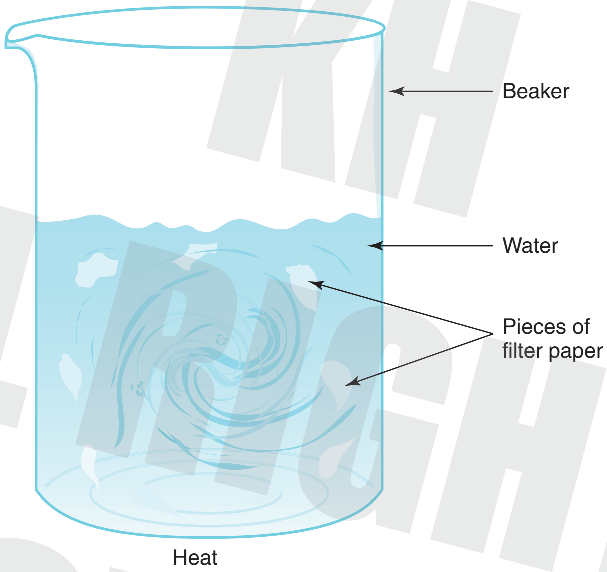

Figure 5.5 Convection currents in water.

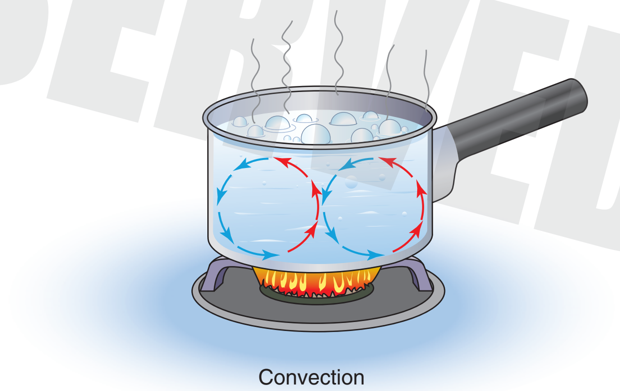

Figure 5.6 Heat Convection.

Remember that convection is responsible for the movement of land and sea breezes. Movements of land and sea breezes are discussed in the Meteorology section. The principle of convection current is evident in our homes. When the fireplace is on, the fire warms the dense cold air, which rises to warm up the house while the cold and dense air sinks to replace the rising warm air, and the process continues.

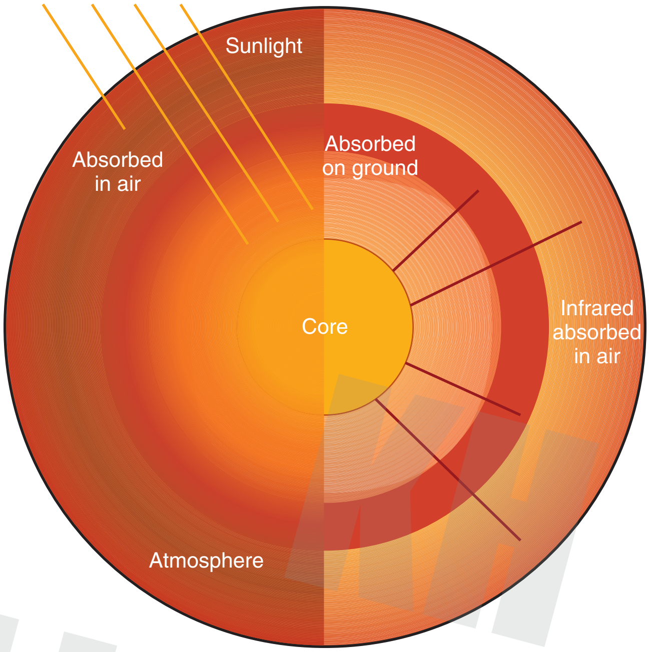

Figure 5.7 Sun's Radiation Reaches the Earth through Space.

5-4-3 Radiation

Radiation of heat, unlike conduction and convection, does not require a material medium. For example, the Sun's energy is transferred to the surface of the Earth by electromagnetic radiation (infrared rays). The radiation travels through space and does not require material medium for heat transfer (Figure 5.7). All materials with temperatures above absolute zero radiate heat (radiant energy). Microwaves in our homes function according to the principle of radiation. You may have wondered why your cold palm during the winter months gets warm when you place it near the electric or gas range. The range releases radiant energy, which is absorbed by your palm, thereby warming it up. Black materials radiate more heat than white materials at the same temperature. Also very good heat radiators are very good absorbers of heat.

5-5 Heat Measurements

Most 18th-century scientists believed heat to be a fluid called *caloric*, flowing from warmer materials to cooler materials. But, an American scientist, Benjamin Thompson, also called Count Rumford (1753–1814), used a metal borer, cannon barrel, and water tank experiment to prove that heat is a form of energy. Rumford's determination was substantiated by James Prescott Joule through his own experiment.

Because heat has been proved to be a form of energy, and the metric unit for energy is joule, therefore, the metric unit for heat is **joule.** It has been experimentally proved by James Prescott Joule that 4.184 joules is equal to one calorie. This is called the *mechanical equivalent of heat*, which indicates that heat and mechanical energy are different ways of representing the same thing. Calorie is also used in the metric system to express heat. Heat is not easily quantified directly. But, variations in temperature are used to quantify heat indirectly. We already know that a rise in temperature means that heat is added, and a fall in temperature shows a loss of heat. **One calorie is the amount of heat required to raise the temperature of one gram of water one degree Celsius**.

One thousand calories is equal to one kilocalorie. One kilocalorie is the amount of heat required to raise the temperature of one kilogram of water one degree Celsius. Nutritionists use *Calorie* to express the amount of energy released when our bodies oxidize food. The Calorie (with a capital *C*) is equal to one kilocalorie. So, do not confuse the calorie with a capital *C* which is a nutritionist's calorie, *C,* and the scientific calorie (cal). The relationship between the two calories is shown below.

1,000 scientific calories = 1 nutritionist Calorie (C)

1,000 scientific calories = 4,184 joules (J)

1 scientific kilocalorie (kcal) = 4,184 joules (J)

You may have wondered how the nutritionists measure the calorific values of food. All they do is to determine the amount of heat released when a specific quantity of food is oxidized. Table 5.2 shows the calorific values of selected food varieties. Calories are per 100 grams of food.

TABLE 5.2 CALORIFIC VALUES OF SELECTED FOOD ITEMS

Food (100 grams)	Approximate Calorie (C)
Cheese	400
Chocolate	548
Egg	167
Lean Meat	286
Liver	143
Milk	71
Peas	100
Sugar	381

The **British thermal unit (Btu)** is an English system of measurement of heat. One Btu is the amount of heat required to raise the temperature of one pound of water one degree Fahrenheit at normal atmospheric pressure. The Btu is used internationally. In the United States it is used to determine the cooling and heating rates of various appliances such as air conditioners. In Europe, commercial gas companies use the Btu unit. One therm is approximately 100,000 Btu or 1.055×10^8 joules (MJ). The calorie, kilocalorie, joules, kilowatts, and Btu are related as shown below.

$$\mathbf{1\ Btu = 252\ calories = 1{,}055\ J = 0.252\ kcal = 2.9 \times 10^{-4}\ kWh}$$

5-6 Specific Heat

Different substances absorb heat energy differently. The differential heat energy absorption depends on the mass and nature of the substance and temperature change. The amount of heat, *Q*, absorbed by a substance is proportional to the mass, *M*, of the substance, temperature change ($T_2 - T_1$) and the nature of the substance (composition of the substance). This can be expressed as

$$\begin{aligned} Q &= cm\,(T_2 - T_1) \\ &= cm\Delta T \end{aligned} \qquad \text{equation 5.4}$$

Q = heat added

where m = mass of the substance in kg

$T_2 - T_1$ = temperature change (ΔT) in °C

c = specific heat capacity of the substance; it is also a proportionality constant which depends on the composition of the substance.

If we solve for c, the equation becomes

$$c = \frac{Q}{m\,(T_2 - T_1)}. \qquad \text{equation 5.5}$$

The unit of specific heat in the metric system **is kcal/kg°C, J/kg°C**or **cal/g°C**. It is **Btu/lb°F** in the English system. The specific heat of a substance is the amount of heat energy needed to raise the temperature of one gram of a substance one Celsius degree. Table 5.3 shows the specific heats of some substances. Specific heats of metals can be determined by a method of mixtures. The popular apparatus is the calorimeter.

TABLE 5.3 SPECIFIC HEAT CAPACITIES OF SOME SUBSTANCES AT ROOM TEMPERATURE AND ATMOSPHERIC PRESSURE UNLESS OTHERWISE STATED

Substance	J/kg/°C or J/kg/K	cal/g/°C or cal/g/K
Water (0°C to 100°C)	4186	1.000
Methyl Alcohol	2549	0.609
Ice (−10°C to 0°C)	2093	0.500
Steam (100°C)	2009	0.480
Benzene	1750	0.418
Wood (typical)	1674	0.400
Soil (typical)	1046	0.250
Air (50°C)	1046	0.250
Aluminum	900	0.215
Marble	858	0.205
Glass (typical)	837	0.200
Iron/Steel	452	0.108
Copper	387	0.0924
Silver	236	0.0564
Mercury	138	0.0330
Gold	130	0.0310

EXAMPLE 5.3

QUESTION:

Calculate the heat required to raise the temperature of 35.0 grams of copper from 25°C to 100°C. The specific heat of copper is 0.0924 cal/g°C.

SOLUTION:

The required heat (Q), mass (m), and change in temperature are related by the following equation.

$$Q = cm\,(T_2 - T_1) = mc\ \Delta T$$

$$m = 35\text{ g}$$

$$T_2 = 100°\text{C}$$

$$T_1 = 25°\text{C}$$

$$C_{copper} = 0.0924\text{ cal/g°C}$$

$$Q = ?$$

Therefore, $Q = (\text{g°C})\ (35\text{ g})\ (100°\text{C} - 25°\text{C})$

$$= \frac{(0.0924\text{ cal/g})\ (35\text{ g})\ (75°\text{C})}{\text{g°C}}$$

$$= \frac{242.6\ \text{cal} \times \text{g} \times °\text{C}}{\text{g°C}}$$

$$= 242.6\text{ cal.}$$

EXAMPLE 5.4

QUESTION:

How much heat must be supplied to 30.0 grams of iron to raise the initial temperature of 20°C to 100°C?

The specific heat of iron is in table 5.3.

Answer: 264 cal.

EXAMPLE 5.5

QUESTION:

You are shopping for a pot that absorbs heat energy fast for a fast cooking time, and you are presented with aluminum and iron pots. In order to make a scientific choice, you are to determine how much heat required to raise the temperature of two 800 g pots from 25°C to 100°C if the pots are made of aluminum and iron respectively.

Answer: Aluminum: 12,900 cal or 12.9 kcal

Iron: 6,600 cal or 6.6 kcal

Your best choice is an iron pot because it absorbs heat energy twice as fast as aluminum.

5-7 Phase Changes

The three states of matter are solid, liquid, and gas. Each state of matter has a different molecular arrangement. The degrees of freedom (motion) of molecules increase drastically from a solid state to a gaseous state. The molecules are most free to move in the gaseous state of matter. Each state changes to another state by either absorption of heat energy or release of heat energy. **Phase change** occurs when one state changes to another state (figure 5.8).

Figure 5.8 shows that below 0°C is the solid state (phase). As heat is applied, temperature increases and the ice molecules separate from the crystal lattice and induce melting. As the ice melts, no amount of heat added will change the temperature. The added heat energy will simply help to change the solid to liquid. This heat energy absorbed by ice without increasing the temperature is called *latent heat of fusion*(L_f).

5-7-1 Latent Heat of Fusion

Latent heat of fusion is **the amount of heat energy added to one gram of a solid at melting stage to change it to a liquid at the same temperature and pressure**. Work must have to be done to separate the molecules of ice from their crystal lattice to induce melting. This work gives the molecules additional potential energy like an object raised to a height acquires potential energy. Remember that the liquid can freeze to become ice. Therefore, the same amount of energy released during melting will be absorbed during freezing. This tells us that the freezing process absorbs energy and the melting process releases energy.

The latent heat of fusion of ice is **80.0 cal/g**. This is the amount of heat energy gram of ice that melts or freezes releases. The more ice, the more the energy is absorbed. In fact, the amount of heat energy needed to change from a solid to a liquid state is the product of the mass of the substance and latent heat of fusion as shown below.

$$Q = mL_f \qquad \textbf{equation 5.6}$$

where L_f is the latent heat of fusion of the solid.

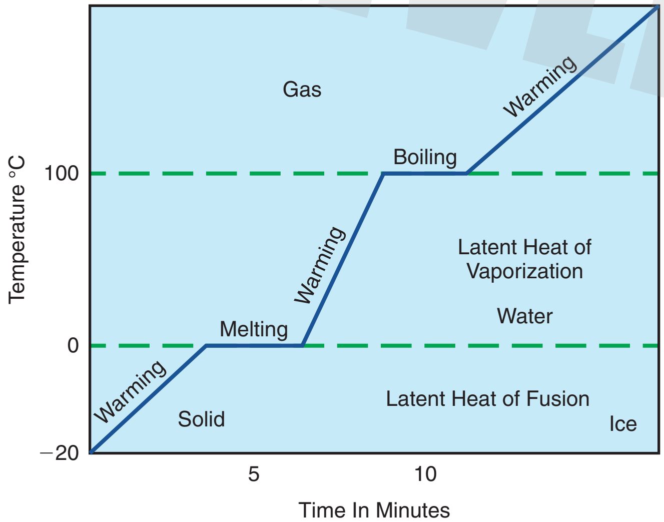

Figure 5.8 Heat Curve.

5-7-2 Latent Heat of Vaporization

Going back to figure 5.8, the heat curve, the **melting point** is the temperature at which the solid changes to liquid. After melting, additional heat energy will increase the temperature to 100°C where the liquid starts to boil. At this point, any additional heat does not change the temperature; instead, it helps to change the liquid to gas. This heat energy absorbed by the liquid without changing the temperature is the **latent heat of vaporization** (L_v) of the liquid. **Latent heat of vaporization** is the amount of heat necessary to change one gram of liquid at its boiling point to gas at the same temperature and pressure. The latent heat of vaporization of water is 540 cal/g. This is the amount of heat energy absorbed by the vaporizing liquid, and the same amount of heat energy must be released to form liquid from gas (condensation). The amount of heat energy necessary to change a liquid to gas is the product of the mass of the liquid and latent heat of vaporization as shown below

$$Q = mL_v \qquad \text{equation 5.7}$$

where L_v is latent heat of vaporization.

5-7-3 Boiling Point

The boiling point is the temperature at which liquid changes to gas. Remember that the temperature increase from one state to another can be determined with the equation

$$Q = mc\,(T_2 - T_1) = mc\Delta T.$$

5-7-4 Sublimation

Sublimation occurs when a solid changes directly to gas without going through the liquid state. An iodine crystal placed in a test tube at room temperature and pressure changes to iodine vapor directly without forming iodine liquid. Also, dry ice (CO_2) placed in a bowl and placed in the open air changes directly to vapor without forming liquid.

EXAMPLE 5.6

QUESTION:

How much heat is required to change 200 g of ice at 0°C to water at 2.0°C?

SOLUTION:

This is a two-step problem. The melting point of ice is 0°C, and the temperature increases from 0°C to form water.

$$Q_1 = mL_f + mc\Delta T \text{ or } mc\,(T_2 - T_1)$$
$$m = 200 \text{ g}$$
$$L_f = 80 \text{ cal/g}$$
$$T = 2.0°\text{C } (2.0°\text{C} - 0°\text{C})$$
$$c = 1 \text{ cal/g°C}$$
$$Q_1 = ?$$

Therefore, $Q_1 = (200 \text{ g})\,(80 \text{ cal/g})$

$$= 16{,}000 \text{ g} \times \text{cal/g}$$
$$= 16{,}000 \text{ cal.}$$

This is the amount of heat energy required to melt 200 g of ice. We can calculate the heat energy required to raise the temperature from 0°C to 2°C.

$$Q_2 = mc\,\Delta T$$
$$= (200\text{ g})\,(1\text{ cal/g°C})\,(2°\text{C})$$
$$= 400\text{ cal}$$

The total latent heat required to change 200 g of ice at 0°C to water at 2.0°C is Q_T.

$$Q_T = Q_1 + Q_2$$
$$Q_T = 16{,}000\text{ cal} + 400\text{ cal}$$
$$= 16{,}400\text{ cal.}$$

EXAMPLE 5.7

QUESTION:

Every household has at least one refrigerator, many hotels have ice makers, and most of us like ice. You may have wondered how much heat energy a refrigerator or ice maker takes from water to make ice. Calculate the amount of energy an ice maker extracts from 1,500 g of water at 25°C to manufacture ice at 0°C.

SOLUTION:

This is a two-step problem.

Step 1. Determine the heat energy removed to cool water from 25°C to 0°C

Step 2. Determine the total heat removed for water to change to ice.

Step 1.

$$Q_1 = cm\,\Delta T$$
$$c = 1\text{ cal/g°C} \qquad m = 1{,}500\text{ g}$$
$$T = 25°\text{C},\ (25°\text{C} - 0°\text{C})$$
$$Q_1 = ?$$

Therefore,

$$Q_1 = (1\text{ cal/g°C})\,(1{,}500\text{ g})\,(25°\text{C})$$
$$= 37500\text{ cal/g°C} \times \text{g} \times °\text{C}$$
$$= 37{,}500\text{ cal.}$$

Step 2.

$$Q_2 = mL_f$$
$$m = 1{,}500\text{ g}$$
$$L_f = 80\text{ cal/g}$$
$$Q_2 = ?$$

Therefore,

$$Q_2 = (1{,}500\text{ g})\,(80\text{ cal/g})$$
$$= 120{,}000\text{ g} \times \text{cal/g}$$
$$= 120{,}000\text{ cal}$$

$$
\begin{aligned}
\textbf{Total heat energy} \quad (Q_T) &= Q_1 + Q_2 \\
Q_T &= 37{,}500 \text{ cal} + 120{,}000 \text{ cal} \\
&= 157{,}500 \text{ cal} \\
&= 157.5 \text{ Kcal} \\
Q_T &= 157{,}500 \text{ cal.}
\end{aligned}
$$

EXAMPLE 5.8

QUESTION:

Calculate the amount of heat required to vaporize 100 g of water at 100°C. The latent heat of vaporization of water is 540 cal/g.

Answer: 54,000 cal.

If two substances of different temperatures undergo heat transformation after the necessary steps have been taken to prevent heat loss to the surroundings, the heat lost by one of the substances will be equal to the heat gained by the other substance. For example, the determination of specific heat of substances with a calorimeter, using the method of mixtures, involves bringing several substances of different temperatures together. When this is done, the hotter substances lose heat and the colder substances gain the lost heat until all the substances reach a common temperature. If no heat is lost to or gained from the surroundings, by the law of conservation of energy, heat lost by the hotter substances is equal to the heat gained by colder substances. This is written heat lost by hotter substances = heat gained by colder substances.

This is expressed with symbols as shown below.

$$Q \text{ lost} = Q \text{ gained}$$

If a metal such as aluminum or copper, water, and a calorimeter are used in the heat transformation, the above expression can be written

$$
\begin{aligned}
Q_{metal} &= Q_{water} + Q_{calorimeter} \\
\text{but} \quad Q &= cm\Delta T \\
\text{Therefore, } (cm\Delta T)_{metal} &= (cm\Delta T)_{water} + (cm\Delta T)_{calorimeter}.
\end{aligned}
$$

EXAMPLE 5.9

QUESTION:

A piece of copper weighing 50 grams is heated to 90°C and placed in a glass of cylinder weighing 100 grams. The glass cylinder contains 150 grams of water at 21°C. The final temperature is 25°C. What is the specific heat of copper? The specific heat of glass is in table 5.3.

SOLUTION:

We will use the following equation:

$$(cm\Delta T)_{copper} = (cm\Delta T)_{water} + (cm\Delta T)_{glass}$$

$$C_{glass} = 0.190 \text{ cal/g°C}$$

$$C_{water} = 1.00 \text{ cal/g°C}$$

$$C_{copper} = ?$$

Therefore,

$$(C)\ (50\text{ g})\ (90°C - 26°C) = (1\text{ cal/g°C})\ (150\text{ g})\ (25°C - 21°C) + (.19\text{ cal/g°C})\ (100\text{ g})\ (25°C - 21°C)$$

$$(C)\ (3{,}250\text{ g°C}) = 600\text{ cal} \times \text{g} \times \text{°C/g} \times \text{°C} + 76\text{ cal} \times \text{g} \times \text{°C/g} \times \text{°C}$$

$$(C)\ (3{,}250\text{ g} \times \text{°C}) = 676\text{ cal}$$

$$(C) = \frac{676\text{ cal}}{3250\text{ g} \times \text{°C}}$$

$$= 0.208\text{ cal/g°C.}$$

5-8 Differences between Heat and Temperature

We often refer to infrared radiation as being primarily heat (or thermal) radiation. But what exactly is heat, and how does it differ from temperature? Simply put, heat is a measurement of energy. All molecules contain some amount of kinetic energy, that is to say, they have some intrinsic motion. The hotter an object is, the faster the motion of the molecules inside it. Thus, the heat of an object is the total energy of all the molecular motion inside that object. Temperature, on the other hand, is a measure of the average heat or thermal energy of the molecules in a substance. When we say an object has a temperature of 100 degrees C, for example, we do not mean that every single molecule has that exact thermal energy. In any substance, molecules are moving with a range of energies, and are interacting with each other as well, which changes their energies. But if we average the thermal energies of all the molecules together, we can obtain the object's temperature.

5-9 Thermodynamics

Thermodynamics deals with transformation of heat to work. Heat engines such as the Carnot, steam turbine, internal combustion, and turbojet engines convert heat into work. In a heat engine, heat is absorbed at a high temperature, which does mechanical work, and eventually released at a lower temperature. Figure 5.9 shows how a heat engine operates. The operation of heat engines is based upon the thermodynamic principles.

The first law of thermodynamics is the same as the law of conservation of energy**. Energy is neither created nor destroyed during the transformation of energy but may be converted**

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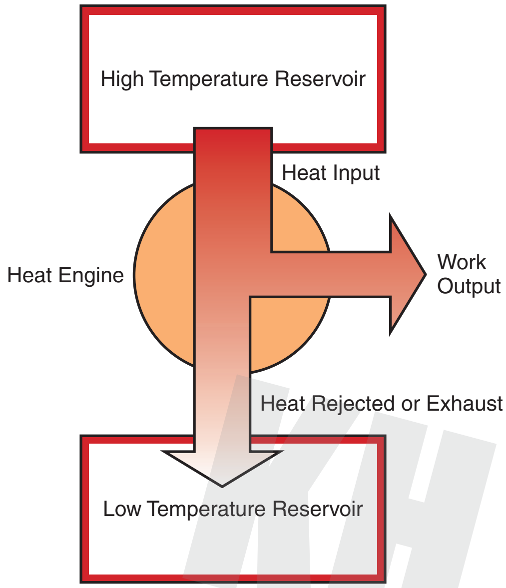

Figure 5.9 A heat engine takes heat from a high-temperature reservoir and changes some of the heat to work and exhausts the rest to a low-temperature reservoir. The work performed by a heat engine is the difference in heat absorbed from high-temperature reservoir and heat released at the low-temperature reservoir.

from one form to another, and the total energy remains the same. Based on the law of conservation of energy, the first law of thermodynamics can be expressed as shown.

Heat input to the system = change in the internal energy of the system + work done by the system.

$$Q = W + E \qquad \textbf{equation 5.8}$$

If there is no heat input or output, Q remains constant. This situation is referred to as *adiabatic process*. When a heat engine operates in a cycle, the internal energy does not change; instead, the net input energy equals the total input energy absorbed by the system from the high-temperature reservoir minus the energy released to the low-temperature reservoir. For example, in an engine that uses gasoline (car engine), at a high temperature the gasoline and oxygen react explosively in each cylinder; and at a low temperature, carbon dioxide or carbon monoxide is released through the exhaust pipe.

All the heat engines mentioned above operate differently from the refrigerator, which is also a heat engine. The heat engines operate by taking in heat from a high-temperature reservoir and releasing it to a low-temperature reservoir with work performed. The refrigerator operates by transforming heat from a low-temperature reservoir to a high-temperature reservoir in the presence of potential energy. You may be thinking that refrigerators generate cold because cold is not generated by heat engines; rather, cold generation in heat engines should be regarded as lack of internal energy. Refrigerators simply transfer internal energies from one point to another.

The second law of thermodynamics states that it is impossible for heat to flow by itself from a low-temperature reservoir to a high-temperature reservoir. This means that it is not possible to design a heat engine that exclusively operates in cycle by absorbing heat from one source and performing equivalent work. Heat normally flows from a high-temperature reservoir to a low-temperature reservoir. It is not possible for water to flow uphill on its own. All rock fragments fall to the foot of a hill forming a talus slope. However, only refrigerators and air conditioners operate in reverse as mentioned earlier.

The second law further shows that an engine that is 100 percent efficient does not exist. A heat engine is not capable of converting all energy input to work. Suppose that the energy input is Q_1, the energy output is Q_2, and the work performed is W, the we have

$$Q_1 = Q_2 + W.$$

This means that not all the energy input is converted into work. Therefore, the efficiency and percent efficiency of any heat engine can be determined as shown below.

$$\text{efficiency (Eff)} = \frac{\text{work}}{\text{energy Input}}$$

$$= W/Q_1. \qquad \text{equation 5.9}$$

$$\text{\% efficiency} = \frac{\text{work}}{\text{energy input}} \times 100$$

$$\text{\% eff} = \frac{W}{Q_1} \times 100 \qquad \text{equation 5.10}$$

$$\text{But} \quad Q_1 = Q_2 + W, \text{ and}$$

$$W = Q_1 - Q_2$$

$$\text{Therefore, Eff} = \frac{Q_1 - Q_2}{Q_1}$$

$$= 1 - Q_2/Q_1$$

$$\text{\% eff} = 1 - Q_2/Q_1 \times 100$$

The maximum efficiency of a heat engine depends on the high temperature reservoir (T_1 in K) and low-temperature reservoir (T_2 in K) and a heat engine cannot exceed its maximum efficiency.

$$Q_1/Q_2 = T_2/T_1$$

because the heat input or output of a Carnot engine is directly proportional to the absolute temperature.

Therefore,

$$\text{Eff} = 1 - \frac{T_2}{T_1}$$

and

$$\text{max. \% eff} = 1 - \frac{T_2}{T_1} \times 100$$

EXAMPLE 5.10

QUESTION:

An engine operates between 590°C and 85°C; calculate a. the maximum efficiency and b. the maximum percent efficiency.

SOLUTION:

Convert the °C to kelvin and use the maximum efficiency and maximum percent efficiency equations.

a.

$$\text{Eff} = 1 - \frac{T_2}{T_1}$$

$$T_2 = 358\text{ K}$$

$$T_1 = 863\text{ K}$$

$$\text{Eff} = ?$$

$$\text{Therefore, Eff} = 1 - \frac{358\text{ K}}{863\text{ K}}$$

$$= 0.56$$

b.

$$\%\ \text{eff} = 1 - \frac{T_2}{T_1} \times 100$$

$$= 1 - \frac{358\text{ K}}{863\text{ K}} \times 100$$

$$= 56\%.$$

EXAMPLE 5.11

QUESTION:

Calculate the maximum efficiency of an efficient heat engine operating between 600°C (T_1) and 96°C (T_2).

Answer: 0.58.

5-10 Entropy

Entropy measures a system's thermal energy per unit temperature that is not available for doing useful work. Entropy was introduced in 1850 by a German mathematician and physicist called Rudolf Clausius (1822–1888). Because work is obtained from ordered molecular motion, the amount of entropy is also a measure of the molecular disorder, or randomness, of a system. Clausius introduced the concept of entropy as a way of expressing the second law of thermodynamics. His form of the second law states that for spontaneous change for an irreversible process in an isolated system (that is, one that does not exchange heat or work with its surroundings) always proceeds in the direction of increasing entropy. For example, a block of ice placed on a hot stove melts, while the stove grows cooler. Such a process is called irreversible because no slight change will cause the melted water to turn back into ice while the stove grows hotter. So, the block of ice and the stove constitute two parts of an isolated system for which total entropy increases as the ice melts. Conversely, a block of ice placed in an ice-water bath will either thaw a little more or freeze a little more, depending on whether a small amount of heat is added to or subtracted from the system. Such a process is reversible because only a very small amount of heat is needed to change its direction from progressive freezing to progressive thawing.

By the Clausius definition, if an amount of heat *Q* flows into a large heat reservoir at temperature *T* above absolute zero, then the entropy increase is $\Delta S = Q/T$. This equation effectively gives an alternate definition of temperature that agrees with the usual definition. Assume that there are two heat reservoirs R_1 and R_2 at temperatures T_1 and T_2 (such as the stoveand the block of ice). If an amount of heat *Q* flows from R_1 to R_2, then the net entropy change for the two reservoirs is $\Delta S = Q\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$, which is positive provided that $T_1 > T_2$. Thus, the observation that heat never

flows spontaneously from cold to hot is equivalent to requiring the net entropy change to be positive for a spontaneous flow of heat. If $T_1 = T_2$, then the reservoirs are in equilibrium, no heat flows, and $\Delta S = 0$.

The condition $\Delta S \geq 0$ determines the maximum possible efficiency of heat engines—that is, systems such as gasoline or steam engines that can do work in a cyclic manner. Suppose a heat engine absorbs heat Q_1 from R_1 and exhausts heat Q_2 to R_2 for each complete cycle. By conservation of energy, the work done per cycle is $W = Q_1 - Q_2$, and the net entropy change is $\Delta S = \frac{Q_2}{T_2} - \frac{Q_1}{T_1}$.

To make W as large as possible, Q_2 should be as small as possible relative to Q_1. However, Q_2 cannot be zero, because this would make ΔS negative and so violates the second law. The smallest possible value of Q_2 corresponds to the condition $\Delta S = 0$, yielding

$$\left(\frac{Q_2}{Q_1}\right)_{\min} = \frac{T_2}{T_1} \qquad \text{equation 5.11}$$

as the fundamental equation limiting the efficiency of all heat engines. A process for which $\Delta S = 0$ is reversible because an infinitesimal change would be sufficient to make the heat engine run backward as a refrigerator.

CHAPTER SUMMARY

Temperature is a measure of the average kinetic energy of the disordered motion of atoms and molecules of a substance. Temperature is measured with a thermometer because a thermometer uses thermometric substances whose physical properties change with temperature change. For example, mercury is a thermometric substance whose physical property changes with temperature change.

Thermometers are designed for measuring temperatures and their construction requires two reference temperatures called lower and upper fixed points or fixed temperatures. **Celsius (centigrade)**, **Fahrenheit,** and **kelvin** are temperature scales. The Celsius scale has an upper fixed point of 100° and a lower fixed point of 0°. The interval between 0° and 100° is called the *fundamental interval*. The Fahrenheit scale has an upper fixed point of 212° and a lower fixed point of 32°. The kelvin scale, also called the absolute scale, has an upper fixed point of 373 K and a lower fixed point of 273 K. No degree sign is attached to a kelvin temperature.

In all the temperature scales, the upper fixed point is the boiling point of water (steam point) and the lower fixed point is the freezing point of water. The Celsius temperature can be converted to the Fahrenheit temperature with the formula $9/5\ Tc + 32°$; and the Fahrenheit temperature can be converted to the Celsius temperature with the formula $5/9\ (T_F - 32°)$. The formula $Tk = Tc + 273$ is used to convert the Celsius temperature to kelvin temperature.

Heat is the measurement of the energy in a substance produced by the molecules in motion. The more energy the molecules of a system possess, the more the heat and higher the temperature. The total energy of the molecules of a system constitute its internal energy. Heat is transferred from an area of higher temperature to an area of lower temperature through conduction, convection, and radiation.

The unit of heat in SI unit is joule, but 4.184 joules equal one calorie. Calorie is also used in SI unit to express heat. One calorie is the amount of heat required to raise the temperature of one gram of water one degree Celsius. The calorie with a capital *C* is the nutritionist's calorie and the calorie with a small letter *c* is the scientific calorie.

The **British thermal Unit (Btu)** is an English system of measurement. The Btu is the amount of heat needed to raise the temperature of one pound of water one degree Fahrenheit, at normal atmospheric pressure.

The **specific heat** of a substance is the amount of heat energy needed to raise the temperature of one gram of a substance one degree Celsius.

Solid, liquid, and gas are the three phases of matter. Phase change occurs when one phase changes to another. **Latent heat of fusion** is the amount of heat energy added to one gram of a solid at melting stage to change it to a liquid at the same temperature and pressure. **Melting point** is the temperature at which a solid changes to a liquid. **Latent heat of vaporization** is the amount of heat necessary to change one gram of a liquid at its boiling point to a gas at the same temperature and pressure. **Boiling point** is the temperature at which a liquid changes to a gas. **Sublimation** is when a solid changes directly from a solid to a gas without going through the liquid phase.

Thermodynamics deals with transformation of heat to work. **The first law of thermodynamics is the same as the law of conservation of energy, and states that energy is neither created nor destroyed but may be converted from one form to another**, and the **total energy remains the same**. The second law of thermodynamics states that it is impossible for heat to flow by itself from a low-temperature reservoir. **The second law further shows that 100 percent efficient engines do not exist because a heat engine is not capable of converting all energy input to work.**

Entropy measures a system's thermal energy per unit temperature that is not available for doing useful work. According to Clausius, if an amount of heat Q flows into a large heat reservoir at temperature T above absolute-zero, then the entropy increase is $\Delta S = Q/T$.

Important Terms

absolute scale
absolute zero
boiling point
British thermal unit
calorie
Celsius scale
centigrade
condensation
conduction
convection
degree Celsius
degree Fahrenheit
efficiency
entropy
Fahrenheit scale
first law of thermodynamics
fixed points
fixed temperatures
freezing point
fundamental interval
gas
heat
heat conductors
heat engine
ice point
insulators
internal energy
joule
kelvin
kelvin scale
kilocalorie
latent heat of fusion
latent heat of vaporization
lattice
liquid
lower fixed point
melting point
molecule
percent efficiency
phases
phase changes
poor heat conductor
solid
specific heat
sublimation
temperature
thermometer
thermometric
vapor

Important Equations

equation 5.1 $T_c = 5/9\,(T_F - 32°)$

equation 5.2 $T_F = 9/5\,T_c + 32°$

equation 5.3 $T_k = T_c + 273$

equation 5.4 amount of heat = (specific heat) (mass) (temperature change)

$Q = cm\,(T_2 - T_1) = cm\Delta T$

equation 5.5 specific heat = amount of heat/(mass) (temperature change)

$C = Q/m\,(T_2 - T_1) = Q/m\Delta T$

equation 5.6 heat input or output = (mass) (latent heat of fusion)

$Q = mL_f$

equation 5.7 heat input or output = (mass) (latent heat of vaporization)

$Q = mL_v$

equation 5.8 first law of thermodynamics: heat input = change in internal energy + work done

$Q = W + WE$

equation 5.9 efficiency = work/energy input

$eff = W/Q_1$

equation 5.10 percent efficiency = work/energy input × 100

$\%\ eff = W/Q_1 \times 100$

MULTIPLE CHOICE QUESTIONS WITH ANSWERS

1. Which of the following is used to measure temperature?
- **a.** anemometer
- **b.** thermometer
- **c.** thermometric substance
- **d.** radiation
- **e.** conduction

2. The thermometric substance used by mercury in a glass thermometer is
- **a.** gas
- **b.** liquid
- **c.** gas and liquid
- **d.** mercury
- **e.** none of the above

3. Mercury may be used as a thermometric substance because
- **a.** the volume of glass increases with temperature
- **b.** the volume of mercury decreases with temperature
- **c.** the volume of glass decreases with temperature
- **d.** the volume of mercury and glass decrease with temperature
- **e.** the volume of mercury increases with temperature

4. Which one of the following does not constitute lower and upper fixed points or reference temperatures in temperature scales?
- **a.** 0° and 100°
- **b.** 32° and 212°
- **c.** 273 K and 373 K
- **d.** 0° and 212°
- **e.** none of the above

5. Which of the following is a fundamental interval?
- **a.** 0° and 212°
- **b.** 273 K and 383 K
- **c.** 0 degrees and 212 degrees
- **d.** 32° and 212°
- **e.** 32° and 100°

6. One of the following units is not used to express temperature:
- **a.** K
- **b.** °C
- **c.** °F
- **d.** °K
- **e.** none of the above

7. The boiling point of water or steam point in the kelvin temperature is
- **a.** 32 K
- **b.** 100 K
- **c.** 212 K
- **d.** 273 K
- **e.** 373 K

8. The freezing or lower fixed point in the Celsius scale is
- **a.** 273°C
- **b.** 32°F
- **c.** 100°C
- **d.** 0°C
- **e.** 273°F

9. The boiling point of water in the Fahrenheit temperature is
- **a.** 212°F
- **b.** 32°F
- **c.** 100°F
- **d.** 0°F
- **e.** 273°F

10. The freezing point of water in the Fahrenheit temperature is
- **a.** 0°F
- **b.** 273°F
- **c.** 373°F
- **d.** 32°F
- **e.** 100°F

11. The boiling point of water in the Celsius temperature is
- **a.** 0°C
- **b.** 32°C
- **c.** 100°C
- **d.** 273°C
- **e.** 373°C

12. The temperature of 10°C equals
- **a.** 50°F
- **b.** 50K
- **c.** 5°F
- **d.** as hot as 2.5°C
- **e.** 32°F

13. The temperature of 60°F equals
- **a.** 15°C
- **b.** 15.8°C
- **c.** 15.56°C
- **d.** 1.5°C
- **e.** 15.66°C

14. The unit of heat in SI units is
- **a.** joule
- **b.** watt
- **c.** newton
- **d.** kg × m/s
- **e.** none of the above

15. One calorie is the amount of heat needed to raise the temperature of one gram of water
- **a.** one degree Fahrenheit
- **b.** one Kelvin
- **c.** one degree Celsius
- **d.** to any temperature
- **e.** one degree Fahrenheit and two degrees Celsius

16. 1,000 scientific calories equal
- **a.** one nutritionist's calorie
- **b.** 486 joules
- **c.** 1,484 joules
- **d.** one megacalorie
- **e.** two kilocalories

17. In the formula $Q = cm\Delta T$, c equals
- **a** Q/m
- **b** $Q/\Delta T$
- **c** $Q-cm\Delta T$
- **d** $Q/m\Delta T$
- **e** Q/T_2-T_1

18. The law that implies that a 100 percent efficient engine does not exist is
- **a.** Newton's first law of motion
- **b.** Newton's second law of motion
- **c.** the first law of thermodynamics
- **d.** the third law of thermodynamics
- **e.** the second law of thermodynamics

19. Which one of the following is a measure of a disorder of a system?
- **a.** entropy
- **b.** temperature
- **c.** heat
- **d.** the first law of thermodynamics
- **e.** adiabatic process

20. Which of the following is not a form of heat transfer?
- **a.** radiation
- **b.** convection
- **c.** conduction
- **d.** sublimation
- **e.** none of the above

21. If two amounts of water have the same temperature, they have the same specific heat, true or false?

22. If two substances of same mass have different specific heat, the same amount of energy is needed to heat them to the same temperature, true or false?

23. Energy transfer from molecule to molecule is called
- **a.** convection
- **b.** radiation
- **c.** conduction
- **d.** equilibrium
- **e.** none of the above

24. Radiation does not require a material medium for heat transfer, true or false?

25. As a liquid goes through a phase change to a gas, heat is absorbed and the temperature
- **a.** increases
- **b.** decreases
- **c.** remains the same
- **d.** fluctuates
- **e.** none of the above

26. Melting point is the temperature at which the solid changes to liquid, true or false?

27. Temperature is not energy, true or false?

28. Temperature is a measure of the average kinetic energy of the disordered motion of atoms and molecules of a substance, true or false?

29. Heat is the measure of the energy in a substance, true or false?

30. Entropy measures a system's thermal energy per unit temperature that is not available for doing useful work, true or false?

31. What is the temperature that is the same numerical value on both the Fahrenheit and the Celsius scale?
- **a.** 212°
- **b.** 0°
- **c.** −40°
- **d.** −273°
- **e.** none of the above

32. The process by which solids changes directly to gas without first becoming a liquid is called
- **a.** sublimation
- **b.** condensation
- **c.** evaporation
- **d.** boiling
- **e.** none of the above

33. This law of thermodynamics shows that an engine with 100% efficiency does not exist.
- **a.** first law of thermodynamics
- **b.** second law of thermodynamics
- **c.** third law of thermodynamics
- **d.** law of increasing entropy
- **e.** none of the above

34. The heat needed to raise the temperature of one gram of water one Celsius degree is
- **a.** joule
- **b.** calorie
- **c.** Btu
- **d.** kilocalorie
- **e.** temperature

35. There is always temperature change during phase changes, true or false?

36. The degrees on the Celsius scale are larger than degrees on the Fahrenheit scale, true or false?

37. The calorie used by nutritionist is kilocalorie, true or false?

38. Which of the following is NOT a way of transferring heat?
- **a.** conduction
- **b.** convection
- **c.** insulation
- **d.** radiation
- **e.** none of the above

39. A material that easily transfers the flow of heat is called
- **a.** collector
- **b.** condenser
- **c.** conductor
- **d.** insulator
- **e.** transfer

40. Iron heats up faster than aluminum does because it has a greater specific heat, true or false?

Answers

1. b, **2.** d, **3.** e, **4.** d, **5.** d, **6.** d, **7.** e, **8.** d, **9.** a, **10.** d, **11.** c, **12.** a, **13.** c, **14.** a, **15.** c, **16.** a, **17.** d, **18.** e, **19.** a, **20.** d, **21.** f, **22.** f, **23.** c, **24.** t, 25. c, **26.** t, **27.** t, **28.** t, **29.** t, **30.** t, **31.** c, **32.** a, **33.** b, **34.** b, **35.** f, **36.** t, **37.** t, **38.** d, **39.** c, **40.** f

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MULTIPLE CHOICE QUESTIONS WITHOUT ANSWERS

1. The thermometric substance used by gas thermometers is
   a. Mercury
   b. Bromine
   c. Germanium
   d. Gas
   e. Cesium
2. Which one of the following is not a temperature scale?
   a. Celsius
   b. Centigrade
   c. Fahrenheit
   d. kelvin
   e. degrees kelvin
3. Which of the following temperature scales is mostly used in weather reports in the United States of America?
   a. Kelvin
   b. Celsius
   c. Fahrenheit
   d. centigrade
   e. all of the above
4. Which of the following is an upper fixed point in a temperature scale?
   a. 32°F
   b. 0°C
   c. 273 kelvin
   d. 212°F
   e. 90°C
5. One of the following pairs is a lower and upper fixed points in temperature scales
   a. 32°F and 200°F
   b. 0°C and 100°C
   c. 273 K and 383 K
   d. 0°C and 212°C
   e. 32°C and 100°C
6. One of the following pairs is not a fundamental interval
   a. 0° and 100°
   b. 32° and 212°F
   c. 273 K and 373 K
   d. 273 K and 383 K
   e. none of the above
7. The temperature of 20°C equals
   a. 68°C
   b. 68K
   c. 6.8°F
   d. 68°F
   e. 680°F
8. The temperature of 100°F equals
   a. 37.9°C
   b. 37.0°C
   c. 36.0°C
   d. 39.7°C
   e. 37.78°C
9. The total energy of the molecules of a system is
   a. internal energy
   b. external energy
   c. a form of energy
   d. fundamental energy
   e. none of the above
10. The law that implies that energy is neither created nor destroyed during transformation of energy, but may be converted from one form to another is
    a. the second law of thermodynamics
    b. the first law of thermodynamics
    c. the third law of thermodynamics
    d. Newton's first law of motion
    e. Newton's second law of motion
11. A teaspoon placed in a hot teacup soon becomes hot because of heat
    a. radiation
    b. convection
    c. conduction
    d. formation
    e. all of the above
12. Which of the following is a heat conductor?
    a. Styrofoam
    b. rubber
    c. wood
    d. glass
    e. iron

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13. In the equation $Q = cm\Delta T$, ΔT equals
- **a.** $Q - cm$
- **b.** $Q \times cm$
- **c.** $Q + cm$
- **d.** Q/cm
- **e.** $Q = cm$

14. One kilocalorie is the amount of heat needed to raise the temperature of one kilogram of water by
- **a.** one degree Fahrenheit
- **b.** one kelvin
- **c.** one degree Celsius
- **d.** one degree Fahrenheit and one degree Celsius
- **e.** none of the above

15. One British thermal unit is the amount of heat required to raise the temperature of one pound of water _______ at normal atmosphere pressure
- **a.** one degree Celsius
- **b.** one degree Fahrenheit
- **c.** one kelvin
- **d.** one kelvin and one Fahrenheit
- **e.** one degree Celsius and two degrees Fahrenheit

16. The amount of heat absorbed by a substance is proportional to
- **a.** the mass of the substance only
- **b.** the mass and temperature change only
- **c.** the temperature change only
- **d.** the temperature and composition of the substance
- **e.** the mass, temperature change, and composition of the substance

17. The amount of heat added to one gram of a solid at melting stage to change it to a liquid at the same temperature and pressure is
- **a.** latent heat of vaporization
- **b.** thermal heat
- **c.** specific heat
- **d.** latent heat of fusion
- **e.** sublimation

18. Melting point is the temperature at which the solid
- **a.** changes to gas
- **b.** sublimates
- **c.** changes to liquid
- **d.** vaporizes
- **e.** boils

19. The change of a solid, liquid, or gas from one state of matter to another is called
- **a.** entropy
- **b.** adiabatic process
- **c.** phase change
- **d.** conduction
- **e.** radiation

20. In absolute temperature scale, the boiling point of water is?
- **a.** 32 K
- **b.** °K
- **c.** 273 K
- **d.** 373 K
- **e.** 373° K

21. If two substances of different mass have different specific heat, the same amount of energy is needed to heat them to the same temperature, true or false?

22. Heat transfer from molecule to molecule is called
- **a.** convection
- **b.** radiation
- **c.** conduction
- **d.** equilibrium
- **e.** none of the above

23. Radiation requires a material medium for heat transfer, true or false?

24. As a solid goes through a phase change to a liquid, heat is absorbed and the temperature
- **a.** increases
- **b.** decreases
- **c.** remains the same
- **d.** fluctuates
- **e.** none of the above

25. The movement of molecules is most restricted in which phase of matter?
- **a.** gas and liquid
- **b.** liquid
- **c.** plasma
- **d.** solid
- **e.** gas

26. Fifteen degree Celsius water is mixed with water at other temperatures to make water at different temperatures; which other water temperature will give the greatest motion to the molecules after they are poured together?
- **a.** 40° C water
- **b.** 60° C water
- **c.** 45° C water
- **d.** 70° C water
- **e.** 80° C water

27. Which phase change occurs at the lowest temperature?
- **a.** sublimation
- **b.** evaporation
- **c.** ionization
- **d.** melting
- **e.** vaporization

28. Which one of the following is not a conductor?
- **a.** air
- **b.** iron
- **c.** aluminum
- **d.** copper
- **e.** silver

29. A water bottle is placed in a refrigerator, and heat is removed from the water. The number of collisions that occur between the water molecules will increase, true or false?

30. If heat is added to a substance, the movement of the molecules and atoms of the substance will
- **a.** stay the same
- **b.** decrease
- **c.** counteract
- **d.** decrease and increase
- **e.** increase

31. The freezing point of a substance is the same as its
- **a.** melting point
- **b.** boiling point
- **c.** condensation point
- **d.** evaporation point
- **e.** none of the above

32. The boiling point of water is 100°Fahrenheit, true or false?

33. If you say that it was so cold yesterday that the temperature only reached 277. Which temperature scale is appropriate?
- **a.** Kelvin
- **b.** Celsius
- **c.** Centigrade
- **d.** Fahrenheit
- **e.** none of the above

34. Water at 23°Fahrenheit is a
- **a.** gas
- **b.** liquid
- **c.** plasma
- **d.** solid
- **e.** none of the above

35. If a weatherman says that the highest temperature for today will be 75 degrees, the temperature scale the weatherman is using is Kelvin, true or false?

36. Through which of the following will convection most likely occur?
- **a.** liquids and gases
- **b.** liquids and solids
- **c.** solids and gases
- **d.** solids and plasmas
- **e.** none of the above

37. Heat and temperature are exactly the same, and they cannot be different, true or false?

38. During a phase change, heat is used to alter the bonding between the molecules of the substance, true or false?

39. Heat change during phase change occurs without change in temperature, true or false?

40. Sun's Radiation reaches the Earth through what?

EXERCISES

Answers and solutions to odd-numbered exercises are in appendix 7.

1. Why is mercury described as a thermometric substance?

2. State the upper and lower fixed points on the temperature scales Celsius, Fahrenheit, and kelvin.

3. If the outside air temperature is 70°C, what is the temperature in Fahrenheit?

4. You are driving to work in the morning and the anchorman announces a temperature of 92°F. What are the Celsius and kelvin equivalents of this temperature?

5. Explain how a microwave functions in terms of heat transfer.

6. Explain the mechanical equivalent of heat.

7. How does a scientific calorie differ from a nutritionist's calorie?

8. **a.** What is the British thermal unit (BTU)?

   b. How many calories and joules are in one BTU?

9. What is specific heat?

10. Determine the heat needed to raise the temperature of 40 grams of copper from 30°C to 100°C. The specific heat of copper is 0.093 cal/g°C.

11. Define a. latent heat of fusion and b. latent heat of vaporization.

12. Determine the amount of heat needed to change 300 g of ice at 0°C to water at 2.0°C.

13. Calculate the amount of energy an icemaker absorbs from 2,000 g of water at 30°C to produce ice at 0°C. Specific heat of water is 1 cal/g°C.

14. How much heat is needed to vaporize 150 g of water at 100°C.

15. If an engine operates between 600°C and 90°C, calculate

a. the maximum efficiency and

b. the maximum percent efficiency of the engine.

16. Determine the maximum efficiency of an efficient heat engine operating between 500°C (T_1) and 95°C (T_2).

17. Explain how a car engine functions in terms of thermodynamic principles.

18. How does a refrigerator function?

19. One degree interval in a Celsius scale equals how many Fahrenheit degree intervals?

20. Three degree intervals in a Fahrenheit scale equals how many Celsius degree intervals?

21. Calculate the amount of heat required to raise the temperature of 450 g of water from 15°C to 85°C? The specific heat capacity of water is 4.18 J/g·°C.

22. If you put 60 grams of ice in your beverage, what quantity of energy would be absorbed by the ice (and released by the beverage) during the melting process? The heat of fusion of water is 333 J/g. Use the formula for latent heat of fusion.

23. Convert 98 degrees Fahrenheit to Kelvin.

24. Calculate the amount of heat needed to change the temperature of 250 grams of ice from 0°C to 2°C?

25. If you like taking bath in warm water and decide to warm 100 kg of water by 20°C, how much heat is required. Give your answer in calories and joules.

26. If an electric current warms up 0.5 kilograms aluminum long to 15°C, calculate the amount of heat produced by the current.

27. What does the temperature scale on the thermometer measure?

28. What happens in the thermometer if the Kelvin temperature of the substance is doubled?

29. Describe what happens if you put a thermometer in your mouth. How does the heat flowing from your mouth affect the atoms in the thermometer?

30. If the air temperature outside changes from morning to noon by about 30°F, what is the corresponding change in temperature in Celsius degrees?

31. If you have a body temperature of 102.5°F, what is your body temperature in degrees Celsius?

32. How much heat must be absorbed by 375 grams of water to raise its temperature by 25°C?

33. What mass of water can be heated from 25°C to 50°C by the addition of 2,825 J?

34. What is the final temperature when 625 grams of water at 75°C loses 7.96 × 104 J?

35. A copper cylinder has a mass of 76.8 g and a specific heat capacity of 0.092 cal/g·°C. It is heated to 86.5°C and then put in 68.7 g of turpentine whose temperature is 19.5°C. The final temperature of the mixture is 31.9°C. What is the specific heat capacity of turpentine?

36. A 65.0 g piece of iron at 525°C is put into 635 grams of water at 15.0°C. What is the final temperature of the water and the iron?

37. If 1 kg ice block at 0°C was placed in a travelling ice box, calculate the amount of heat released by the ice as it melts to water.

38. What is the difference between latent heat of vaporization and latent heat of fusion?

39. List all phase changes.

40. Describe how heat behaves during sublimation (solid to gas).