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***Physics*, Chapter 15: Heat and Work**

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15

Heat and Work

15-1 The Nature of Heat

Until about 1750 the concepts of heat and temperature were not clearly distinguished. The two concepts were thought to be equivalent in the sense that bodies at equal temperatures were thought to “contain” equal amounts of heat. Joseph Black (1728–1799) was the first to make a clear distinction between heat and temperature. Black believed that heat was a form of matter, which subsequently came to be called *caloric*, and that the change in temperature of a body when caloric was added to it was associated with a property of the body which he called the capacity. Later investigators endowed caloric with additional properties. The caloric fluid was thought to embody a kind of universal repulsive force. When added to a body, the repulsive force of the caloric fluid caused the body to expand. To explain the liberation of heat when a block of metal was filed, it was postulated that small filings were less able to retain caloric, by virtue of their large surface area, than a block of solid metal. Attempts were made to measure the weight of caloric by trying to observe a change in the weight of a body when its temperature was raised, but these experiments were contradictory. Among others, Count Rumford (1753–1814), an American born Benjamin Thompson, who gained his title in the service of the Elector of Bavaria, found that the weight of a block of gold was unaltered by as much as 1 part in 1,000,000 when raised from the freezing point of water to bright-red heat.

As director of the arsenal at Munich, Rumford observed that the heat generated in boring cannon did not depend upon the number of chips evolved. He performed experiments which showed that the heat capacity of brass chips was the same as that of solid brass, contradicting the theory advocated by the adherents of the caloric theory of heat. He concluded that the evolution of heat during the operation of boring a cannon could not be attributed to the liberation of caloric, and he set up an experiment in which the heat evolved in rubbing a blunt horse-driven drill against a

brass casting was used to heat a known quantity of water. Rumford was able to trace the heat evolved to the effort of the horse rather than to the ability of brass filings to retain caloric fluid.

It was not until the middle of the nineteenth century that the presently accepted view of the nature of heat evolved. In 1842 J. R. Mayer laid the foundation of the principle of the conservation of energy, proposing that heat was a form of energy. He utilized some well-known data on the heat necessary to maintain the temperature of expanding air, which normally cools on expansion, to compute a numerical relationship between units of heat and units of mechanical energy which is called the *mechanical equivalent of heat*. The suggestion of Mayer was investigated in great detail by James Prescott Joule (1818–1889), who measured the mechanical equivalent of heat by every experimental means at his disposal. Joule utilized mechanical energy to churn water and mercury. He generated heat frictionally and electrically. The experiments of Joule firmly established the concept that heat was a form of energy and that an equivalent quantity of heat was always obtained when a given quantity of mechanical energy was expended.

We have already shown that temperature is a property of a system which determines whether or not it is in equilibrium with any other system placed in thermal contact with it. If two systems which are at different temperatures are placed in thermal contact, energy will flow from the one at higher temperature to the one at lower temperature. We shall use the term *heat* to designate the energy in transit owing to a difference in temperature between two systems. The details of the transformations which occur when this energy enters or leaves a system form an important part of physics, and we shall consider many of these transformations in this book. Some of these will be considered from the macroscopic or large-scale point of view, while other transformations will be considered from the microscopic point of view, involving some ideas concerning the structure of matter.

We have already outlined briefly (Section 13-1) our present view concerning the structure of matter with particular reference to the forces existing between molecules. Another approach to the subject of the structure of matter is in terms of the energy of the constituent particles, not only of the molecules of the substance but also of their atoms and of the electrons, protons, and neutrons which constitute these atoms. Each of these constituent particles undoubtedly has energy, probably both kinetic and potential. A complete theory of the structure of matter would enable us to evaluate the energy of each particle under all possible conditions, and from this we would then be able to evaluate the energy of the substance under any given set of physical conditions. This energy constitutes the *internal energy* of the substance under these conditions. However, there is

at present no satisfactory theory which will enable us to evaluate the internal energy of a substance. There is, however, another approach by which we can evaluate *changes in the internal energy* of a substance. For example, if a known quantity of heat is added to a substance and no external work is done on or by this substance, then this quantity of heat will be converted into additional internal energy of the substance. The complete theory of the structure of matter would be able to trace the changes in the energies of the individual particles of the substance corresponding to the quantity of heat added. However, the theory of the structure of matter in its present state is unable to provide us with a detailed analysis of how the internal energy of a substance is distributed among its constituent particles. But we can say that if U_i is the internal energy of the substance before a quantity of heat Q is added to it, and if U_f is the internal energy of the substance after the addition of this amount of heat, then

$$Q = U_f - U_i, \quad (15-1)$$

providing no external work was done on or by the substance.

One might inquire as to what evidence we have that the internal energy has been changed by the addition of heat to a substance. The evidence is of an indirect kind. We may note that the temperature of the substance has been changed, or that its physical state has been changed—say from a liquid to a gas—or that its pressure has been changed. A complete theory would relate the energies of the particles of the substance to its temperature, pressure, and physical state. To date, we have made only small progress in this direction. Probably the greatest progress has been made in the study of the behavior of substances which approximate closely the behavior of *perfect* or *ideal gases*. This will be studied in greater detail in the next chapter. We may anticipate this by stating that if the assumption is made that the absolute temperature of a perfect gas is proportional to the average kinetic energy of the gas molecules, the results are consistent with the known properties of a gas. While this direct proportionality between temperature and the kinetic energy of molecules holds for an ideal gas, it does not necessarily follow that it holds for a real substance in the solid, liquid, or gaseous phase. For real substances, the relationship between temperature and the kinetic energy of the molecules is a more complicated one.

When a body is translated with uniform speed there is an ordered motion of the body and it has kinetic energy, but there is no alteration of the internal energy of the body that is associated with the random motions of the fundamental particles composing the body. Consequently the temperature of a body is unaltered by a uniform translational motion.

15-2 Units of Heat; Heat Capacity

Since heat is a form of energy, any of the energy units such as the erg, joule, or foot pound may be used to measure a quantity of heat. However, it is usually more convenient to use some other unit of heat defined in terms of one of the effects produced by the addition of heat to a body, and then to determine the relationship between this heat unit and the more common energy units.

By tradition, heat is measured in terms of the temperature changes produced in water. Several units of heat are in common use. The *calorie* (abbreviated cal) is defined as *the quantity of heat which will raise the temperature of one gram of water from 14.5°C to 15.5°C*. This unit of heat is called *the 15-degree calorie*. In biology, in dietetics, and in some branches of engineering, a unit of 1,000 calories is frequently used. This is called the *large calorie*, or *kilogram calorie* (abbreviated kilo cal), for it is *the heat required to raise one kilogram of water through the above temperature interval*. The *British thermal unit* (abbreviated Btu) is defined as *the quantity of heat which will raise the temperature of a pound mass of water from 63°F to 64°F*.

The quantity of heat that is required to raise the temperature of a body by one degree is called its *heat capacity*. The heat capacity of a unit mass of a substance is called its *specific heat*, and is usually designated by the symbol c . The specific heat is generally stated in units of calories per gram per degree centigrade, or in units of Btu per pound per degree Fahrenheit, and has the same numerical value in both systems of units.

As shown in Figure 15-1, the specific heat of water in the interval 0°C to 100°C is not constant. The variation is of the order of $\frac{1}{2}$ per cent and will be neglected in the calculations performed in this chapter. Under these circumstances we may abridge the definitions of the calorie and the Btu, so that for practical purposes these are, respectively, the heat required to raise the temperature of a gram of water one degree centigrade, and the heat required to raise the temperature of a pound of water one degree Fahrenheit. Since there are approximately 454 gm in a pound mass, a Btu of heat will

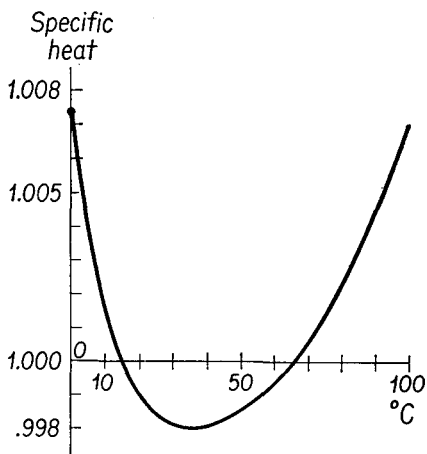


Fig. 15-1 Specific heat of air-free water as a function of temperature at 1 atm pressure.

raise the temperature of 454 gm of water by $\frac{5}{9}^{\circ}\text{C}$, and is therefore approximately equal to 252 cal.

From the definitions of heat capacity and specific heat, it is evident that the heat capacity of a body may be expressed as the product of the mass of the body by the specific heat of the substance from which it is made. It is also common practice to refer to the *water equivalent* of a body as the mass of water numerically equal to the heat capacity of the body.

When a body of mass m is heated through a temperature interval Δt , the heat Q required is given by

$$Q = mc \Delta t, \quad (15-2)$$

where c is the specific heat of the body. The specific heats of some common substances are given in Table 15-1. In general, the specific heat varies

TABLE 15-1 SPECIFIC HEATS

Substance	Specific Heat	
	$\frac{\text{cal}}{\text{gm}^{\circ}\text{C}}$	or $\frac{\text{Btu}}{\text{lb mass } ^{\circ}\text{F}}$
Aluminum	(-250°C)	0.0039
	(-100°C)	0.1676
	(20°C)	0.214
	(100°C)	0.225
Brass	($20-100^{\circ}\text{C}$)	0.092
Copper	($15-100^{\circ}\text{C}$)	0.093
Crown glass	($10-50^{\circ}\text{C}$)	0.161
Ice	(-20°C)	0.48
	(-10°C)	0.53
Steam, 1 atm	(110°C)	0.481
	(120°C)	0.477
Iron	(20°C)	0.107
Lead	(20°C)	0.031
Mercury	(20°C)	0.033
Silver	(20°C)	0.056
Tin	(20°C)	0.054
Zinc	(20°C)	0.092
Ammonia (liquid)	(20°C)	1.12
Ethyl alcohol	(25°C)	0.456

with temperature; the heat required to raise a body from some initial temperature t_i to a final temperature t_f can be found by integration if c is

known as a function of temperature. Thus

$$Q = \int_{t_i}^{t_f} mc dt.$$

For small temperature intervals, and for the calculations of this chapter, we shall neglect the variations of the specific heat with temperature.

The heat capacity of a mole of a substance is often more interesting theoretically than the heat capacity of a unit mass. The heat capacity of a mole is referred to as the *molar heat capacity* of the substance. This is given by the product of the specific heat by the molecular weight. It is interesting that, over a century ago, Dulong and Petit observed that the product of the specific heat by the atomic weight of many crystalline solids, composed of elements rather than compounds, was constant. The average atomic heat capacity of some 63 elements is about 6.15 cal/gm atomic wt, while the individual values of these substances ranged between 5.4 and 7.0 cal. Modern research has shown that these results depend on the temperature, for the specific heat of a solid approaches zero as the temperature approaches the absolute zero. There are similar generalizations which can be made regarding the molar heat capacity of gases. We shall discuss some of the implications of these generalizations in a subsequent chapter.

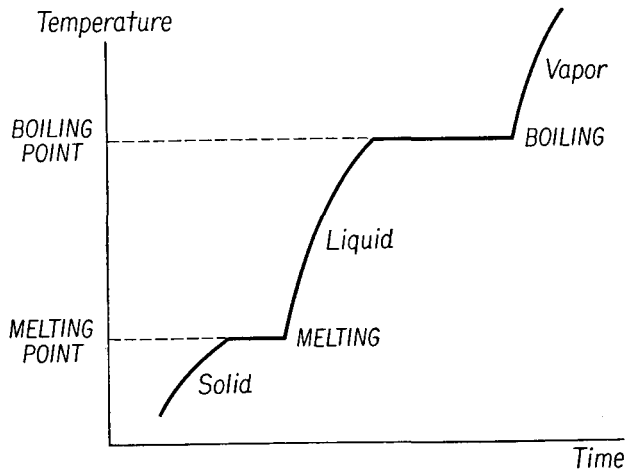


Fig. 15-2

15-3 Latent Heat

When a solid substance is heated at a uniform rate, its temperature rises steadily until its melting temperature is reached, as shown in Figure 15-2, and the substance becomes liquid. During the interval of time that the

temperature of the substance remains constant, heat is being continually absorbed. The substance is melting, and it coexists in the solid and liquid phases. In general, when the phase of a substance is altered, as from solid to liquid or from liquid to vapor, or when the crystalline form of the solid is altered, the internal energies of the two phases are different, and heat is absorbed by the substance making the phase change. An equal amount of heat is liberated when the phase change is made in the reverse direction. This heat is called the *latent heat*. The heat required to change a unit mass of substance from the solid to liquid phase at constant temperature is called the *latent heat of fusion*. If we continue to heat the substance shown in Figure 15-2, the temperature of the liquid continues to rise steadily until the boiling point is reached. Once again the rise in temperature is arrested until the liquid is completely converted to a vapor. The heat required to change a unit mass of liquid from the liquid to the vapor phase at constant temperature is called the *latent heat of vaporization*. As we will see in a subsequent chapter, the temperature at which the phase change occurs is a function of the pressure, and the latent heat associated with a particular phase change is also dependent upon the pressure and temperature at which it occurs.

For the water substance at standard atmospheric pressure, the latent heat of fusion is approximately 80 cal/gm. The latent heat of vaporization at standard atmospheric pressure is approximately 540 cal/gm. The latent heat, generally represented by the symbol L , is associated in part with the differences in internal energy of the two phases and in part with the mechanical work which must be done by the substance in pushing back the atmosphere.

15-4 Calorimetry

One of the methods used in heat measurements, or *calorimetry*, is called the *method of mixtures*. In this method a hot body at some temperature t_1 is placed in thermal contact with a second body at temperature t_2 . There is a heat exchange between the two bodies, and, when equilibrium is reached at some intermediate temperature t , *the heat gained by the cooler body must be equal to the heat lost by the warmer body*, assuming that there have been no losses of heat to the outside world. One way to minimize heat losses is to arrange the experiment so that the final equilibrium temperature is approximately equal to the ambient temperature of the laboratory. Another way is to surround the bodies with an enclosure which is maintained at the average temperature of the system with external sources of heat and thermostatic controls.

Illustrative Example. Twenty grams of ice, originally at -10°C , are put into a copper can of mass 60 gm containing 180 gm of water at 80°C . The copper

can is called a *calorimeter*. Find the final equilibrium temperature of the system. Assume that we may neglect the loss of heat to the outside world.

Let us assume that in the final equilibrium condition all the ice has melted, and that equilibrium temperature of the system is t . To reach this state the temperature of the ice must first be raised to 0°C , and in this process the ice must gain an amount of heat given by

$$Q_1 = m_i c_i \Delta t_i = 20 \text{ gm} \times 0.5 \frac{\text{cal}}{\text{gm}^\circ\text{C}} \times 10^\circ\text{C} = 100 \text{ cal}.$$

At 0°C the ice must acquire its latent heat of fusion to melt to water, and the heat gained in this process is given by

$$Q_2 = m_i L = 20 \text{ gm} \times 80 \frac{\text{cal}}{\text{gm}} = 1,600 \text{ cal}.$$

Finally the ice water must be raised to a final temperature t , gaining heat in amount given by

$$Q_3 = m_i c_w \Delta t = 20 \text{ gm} \times 1 \frac{\text{cal}}{\text{gm}^\circ\text{C}} \times (t - 0^\circ\text{C}) = 20 t \frac{\text{cal}}{^\circ\text{C}}.$$

The total heat gained by the ice in reaching its final equilibrium state as water at temperature t is the sum of Q_1 , Q_2 , and Q_3 , and is given by

$$Q_1 + Q_2 + Q_3 = 1,700 \text{ cal} + 20 t \frac{\text{cal}}{^\circ\text{C}}.$$

At the same time the calorimeter can and the water must lose an equal amount of heat in being cooled from 80°C to t . The heat lost by the copper calorimeter can is given by

$$Q_4 = m_{\text{cu}} c_{\text{cu}} \Delta t = 60 \text{ gm} \times 0.093 \frac{\text{cal}}{\text{gm}^\circ\text{C}} \times (80^\circ\text{C} - t),$$

or
$$Q_4 = 447 \text{ cal} - 5.6 t \frac{\text{cal}}{^\circ\text{C}}.$$

The heat lost by the water in cooling from 80°C to t is

$$Q_5 = m_w c_w \Delta t = 180 \text{ gm} \times 1 \frac{\text{cal}}{\text{gm}^\circ\text{C}} \times (80^\circ\text{C} - t),$$

or
$$Q_5 = 14,400 \text{ cal} - 180 t \frac{\text{cal}}{^\circ\text{C}}.$$

The total heat lost by the water and the calorimeter can is

$$Q_4 + Q_5 = 14,847 \text{ cal} - 185.6 t \frac{\text{cal}}{^\circ\text{C}}.$$

Finally we solve the problem by equating the heat gained to the heat lost,

$$1,700 \text{ cal} + 20 t \frac{\text{cal}}{^\circ\text{C}} = 14,847 \text{ cal} - 185.6 t \frac{\text{cal}}{^\circ\text{C}},$$

$$205.6 t \frac{\text{cal}}{^\circ\text{C}} = 16,547 \text{ cal},$$

$$t = 8.1^\circ\text{C}.$$

Thus the final temperature of the mixture is 8.1°C .

A second calorimetric procedure, based upon the melting of ice at 0°C , is called the *ice calorimeter*. A block of ice at 0°C is carefully dried, and provided with a cavity and an ice cover. A heated specimen at known initial temperature t is placed in the cavity. When the specimen has cooled to the ice point, the water in the cavity is collected and weighed. Since the water is at a temperature of 0°C , the only heat gained by the ice which has melted is associated with the latent heat of fusion. We may write

$$m_w L_f = mc(t - 0),$$

where m_w is the mass of the water in the cavity, and m , c , and t are the mass, specific heat, and initial temperature of the specimen. The ice calorimeter was used by Bunsen to determine the specific heat of many metals. There is little difficulty with loss of heat from the interior cavity of the ice calorimeter, for heat exchanges with the laboratory affect the outside surface of the block rather than the interior cavity. The greatest difficulty lies in completely collecting the water which has been melted.

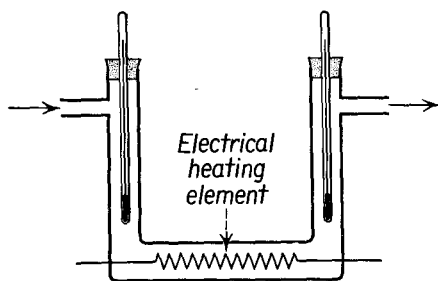


Fig. 15-3 Schematic diagram of a continuous-flow calorimeter.

The *heat of combustion* of a solid fuel, or of a food substance, may be determined in a *bomb calorimeter*. A known mass of the substance is placed in a strong container, or bomb, which has been provided with a supply of oxygen, often in chemical form, and a means of igniting the mixture. The bomb is placed in water, and the heat liberated in the combustion may be calculated by the principles of the method of mixtures. Similar procedures may be used to measure the heat of formation and the heat of solution of chemical compounds.

In the *continuous-flow calorimeter*, shown in Figure 15-3, a steady source of heat is supplied to a fluid flowing through the calorimeter. The rate of flow and the temperature of the fluid as it passes into and out of the calorim-

eter are measured. The heat of combustion of a fuel may be measured by burning the fuel in the calorimeter. If it is desired to measure the specific heat of a liquid, it is heated by an electrical heater which evolves heat at a known rate; measurement of the rate of flow of the liquid and the intake and outlet temperatures, provide the data required to calculate the specific heat.

It should be noted that the specific heat measured and used in most calorimetric experiments is an average specific heat in which the average is taken over the temperature interval used in the experiment. It is much more difficult to measure the specific heat at a particular temperature, and the methods for this measurement will not be elaborated here.

In all heat measurements it is necessary to make rather elaborate corrections for the interchange of heat with the laboratory, and to make suitable provision for good thermal contact. One cannot properly measure the temperature of a block of metal simply by placing a thermometer upon it. It is necessary to pay attention to the uniformity of temperature of the specimen, to the thermal contact, and to thermal losses from the measuring device.

15-5 Work and Heat

One form of the apparatus used by Joule in his experiments to determine the mechanical equivalent of heat is illustrated in Figure 15-4. The system on which work was done was a quantity of water of mass M contained in a well-insulated glass jar of mass M' . A set of paddles A was attached to a shaft which was set in rotation by a set of weights W , connected to the paddles through a system of strings and pulleys. The motion of the paddles A past the stationary vanes B churned the water, thus increasing its temperature through a temperature interval Δt . If the weights started from rest and acquired a velocity v after falling through a height h , the mechanical energy \mathcal{W} lost by the weights and delivered to the water was

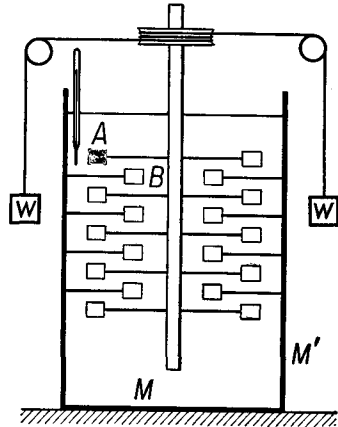


Fig. 15-4 Joule's apparatus for determining the mechanical equivalent of heat.

$$\mathcal{W} = mgh - \frac{1}{2}mv^2, \quad (15-3)$$

where m is the total mass of the falling weights. The amount of heat that would have to be added to the water (and the jar) to produce the same

change in temperature is

$$Q = Mc_w \Delta t + M'c_g \Delta t. \quad (15-4)$$

Equation (15-3) gives the work done expressed in mechanical energy units, while Equation (15-4) gives the equivalent amount of heat in heat units. From this and similar experiments it is found that

$$4.185 \times 10^7 \text{ ergs} = 1 \text{ cal},$$

$$4.185 \text{ joules} = 1 \text{ cal},$$

$$778 \text{ ft lb} = 1 \text{ Btu}.$$

If we desire to equate the values of \mathcal{W} and Q in an equation, it is necessary to introduce a conversion factor J representing the mechanical equivalent of heat. Thus

$$\mathcal{W} = JQ, \quad (15-5)$$

where \mathcal{W} is the energy expressed in mechanical units and Q is the energy expressed in heat units. The value of J may then be expressed in appropriate units as

$$J = 4.185 \text{ joules/cal},$$

or

$$J = 778 \text{ ft lb/Btu}.$$

15-6 The First Law of Thermodynamics

We are now in a position to extend the principle of conservation of energy to include heat energy and internal energy as well as mechanical energy. When so formulated, the principle of conservation of energy is called the *first law of thermodynamics*. Let us consider that the substance or device under investigation is within an imaginary closed shell. We shall call everything within that shell the *system*. The shell may simply enclose a block of metal, or it may be so large as to enclose the entire solar system.

To develop our ideas about a suitable sign convention, let us pattern our thinking after the operation of an engine in which it is desired that the engine absorb heat from some source and do mechanical work. A quantity of heat ΔQ delivered to the system will be called positive. If heat is expelled by the system, we shall call that heat negative. A quantity of work $\Delta \mathcal{W}$ done by the system on the outside world will be positive work. Mechanical work done by the outside world upon the system will be negative. In general, the result of adding a quantity of heat ΔQ to a system is that an amount of work $\Delta \mathcal{W}$ will be done by the system on some outside agency, and the internal energy of the system will be changed by an amount ΔU .

The first law of thermodynamics can be written in the form of an equation as

$$\Delta Q = \Delta U + \Delta \mathcal{W}, \quad (15-6)$$

where all quantities must be expressed in the same units. Unlike the usage of the suffix *dynamics* in the word “hydrodynamics,” the word “thermodynamics” does not refer to heat in motion. The subject of thermodynamics deals with the relationship between heat and mechanical energy, as typified in Equation (15-6).

If no work is done on or by the system, the change in its internal energy is equal to the heat added to or removed from the system. The addition of heat to a system increases its internal energy. Evidence for the increase in internal energy is often the increase in the temperature of the system.

A physical process in which no heat enters or leaves the system is called an *adiabatic* process. In an adiabatic process $\Delta Q = 0$. Many of the processes of nature and of engineering are adiabatic or nearly so. A process which takes place inside a chamber with perfectly insulating walls is adiabatic. Many processes which take place quickly, so that there is little time for heat to be transferred, are nearly adiabatic. Thus, when air is compressed in a bicycle pump, the compression is practically adiabatic.

An interesting aspect of the first law of thermodynamics is that if we grant the impossibility of constructing a perfectly frictionless machine, Equation (15-6) asserts the impossibility of constructing a machine which will run forever; that is, a special type of perpetual-motion machine. If no mechanical work is done on or by the machine, the quantity $\Delta \mathcal{W}$ is equal to zero. Since heat is generated by friction in the bearings and is expelled from the system, the quantity ΔQ is negative. Thus the internal energy of the system must decrease, and the machine must ultimately run down.

Problems

15-1. An aluminum calorimeter whose mass is 100 gm contains 200 gm of water at 24°C. A mass of 500 gm of tin at a temperature of 95°C is placed in the water. Determine the final temperature of the mixture.

15-2. In determining the specific heat of a metal, a 150-gm sample is heated to 99.5°C and then put into 225 gm of water at 18°C which is contained in a copper calorimeter whose mass is 160 gm. The final temperature of the mixture is 22.4°C. Determine the specific heat of the sample.

15-3. A metal calorimeter whose heat capacity is 20 cal/°C contains 300 gm of water at 10°C. One hundred grams of copper at 150°C and 250 gm of zinc at 115°C are placed in the water. What is the final temperature of the mixture?

15-4. A lead ball weighing 1 lb is dropped from a height of 80 ft to the pave-

ment. Assuming that all of this energy is converted into internal energy of the lead ball, determine the rise in temperature of the ball.

15-5. The water at Niagara Falls drops 160 ft. Assuming that all of the energy of the water is converted into internal energy, what is the increase in temperature of the water after it strikes the bottom of the falls?

15-6. The powder which is used to fire an 8-gm bullet produces 800 cal of heat when burning. In the firing of the bullet, only 25 per cent of this energy is converted into kinetic energy of the bullet. Determine the muzzle velocity of the bullet.

15-7. A large glass jar whose mass is 250 gm contains 2,000 gm of water at 18°C. An electric motor delivers $\frac{1}{8}$ hp to a stirrer in the water for 10 min. Determine the rise in temperature of the water if all this energy goes into the water and the jar.

15-8. A copper cylinder whose mass is 50 gm contains 200 gm of water. The cylinder is dragged over a rough floor by a force of 100,000 dynes, and the acceleration produced is observed to be 200 cm/sec². (a) Determine the force of friction between the cylinder and the floor. (b) If the cylinder is moved through a distance of 100 m, determine the rise in temperature of the water and the cylinder, assuming that all of the work against friction is converted into internal energy in the cylinder.

15-9. A block of 500 gm of copper at 750°C is dropped into an ice calorimeter. How many grams of ice are melted?

15-10. A container, of water equivalent 5 gm, contains 45 gm of ice at -10°C. What will be the final condition and temperature of the material in the container if 100 gm of water at 20°C are poured into the container?

15-11. An ice crusher containing 1000 gm of ice at 0°C is driven for 15 min by a 100-watt motor. What is the final temperature of the ice, assuming that all of the energy of the motor is converted to internal energy of the ice? How much of the ice will be melted?

15-12. A container of crushed ice is at -15°C. The water equivalent of the container is 20 gm, and the ice has a mass of 200 gm. How much steam at 100°C has to be passed into the container in order that the ice be melted and the entire system, consisting of ice water, condensed steam, and container, reach a final temperature of 20°C?

15-13. A tank of water containing 100 lb of water is to be heated from 50°F to 65°F in 15 min by an electric heater. If the container is of negligible heat capacity, what is the minimum rating, in watts, of the heater which should be purchased to do this job?

15-14. A liter of water per minute flows through a tank which is heated by an electric heating element. The water flows into the tank at a temperature of 55°F and is to flow out of the tank at a temperature of 155°F. What should be the minimum rating, in watts, of the heater required to do this job?

15-15. An electric mixer is supplied with electrical energy at the rate of $\frac{1}{8}$ hp and delivers mechanical energy to a bowl of batter at the rate of 5,000 joules/min. (a) What is the efficiency of the mixer? (b) What is the increase in the internal energy of the motor in 10 min? (c) If the water equivalent of the motor is 1,000 gm, what is the increase in the temperature of the motor in 10 min?