Chapter 10: Temperature and Energy

Have you ever burned your tongue on hot apple pie? If so, you have probably also noticed that, while you must be careful about the apple filling, the crust is not dangerous even when the pie has just come out of the oven. Visitors to the San Francisco Bay region know the remarkably mild climate enjoyed by the environs of the Bay, where the summer temperature is often 20 or more degrees less than it is only a few miles inland. Both the apple pie effect and the San Francisco climate are related to the fact that water is different from most other materials in its capacity to store thermal energy. All substances become colder or hotter as they lose or gain thermal energy (assuming they are not changing phase, such as in melting or boiling). For water, however, the temperature change is relatively small for a given energy change. The wet pie filling, therefore, transfers energy to the tongue and still remains hot, while the crust cools off substantially when it interacts with the tongue. Conversely, both land and water receive similar amounts of energy from the sun in the summer, but San Francisco Bay water rises in temperature much less than does the adjacent land.

These two explanations rest on your awareness that thermal energy and temperature are distinct factors, though they usually increase or decrease together. The definition of the Calorie in Section 9.2 makes it possible for you to measure the thermal energy stored in any system in which you may be interested. We still must explain the measurement of temperature. In this chapter, therefore, we will explore the relationship of thermal energy to the temperature of a system and make mathematical models for the relationship. By using these models and thought experiments, you will be able to calculate the equilibrium temperature of a system whose subsystems were initially at different temperatures.

Chemical energy, phase energy (the energy of liquid versus solid or gas versus liquid at the same temperature), and thermal energy are easily transformed into one another, as you know from burning a candle and making iced tea. We will therefore include all three forms of energy storage in our discussion. Finally, we will describe how the many-interacting-particles (MIP) model for matter explains the thermal energy transfer that is observable in the macro domain.

10.1 Operational definition of temperature

An instrument for measuring temperature is called a *thermometer.* You can make a thermometer out of any system that gives easily observed evidence of a change in its temperature (Fig. 10.1). Such a system is called a *thermometric element.* Thermal expansion of solids or liquids is the property of matter most commonly used in the construction of thermometers, but Galileo used air in his "thermoscope" (Fig. 10.2). Since the expansion effects are small, one practical procedure is to confine an amount of liquid (mercury, red-dyed alcohol, or kerosene) in a bulb and to let the liquid extend into a very fine tube. Even a small change in volume brought about by warming or cooling of the liquid will then result in a visible change in length of the liquid in the tube
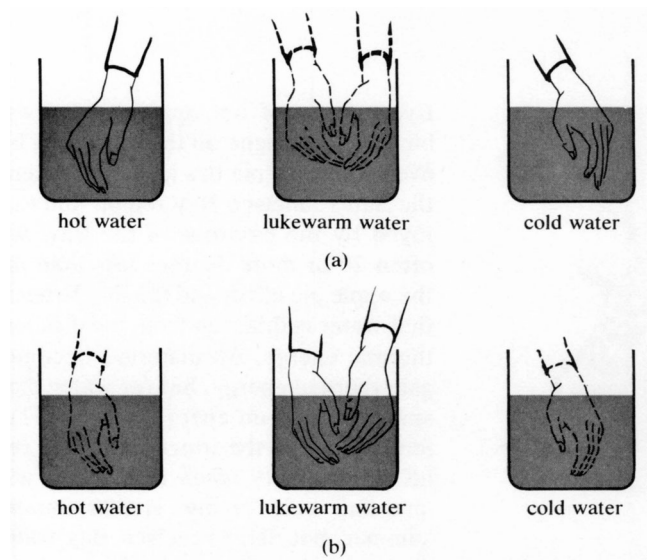
249

hot water   lukewarm water   cold water

(a)

hot water   lukewarm water   cold water

(b)

Figure 10.1 (above) Is your hand a good thermometer? To find out, proceed as follows:
(1) Place both hands in the lukewarm water. Do both hands sense the same temperature?
(2) Place one hand in the hot water, and one in cold, for a minute.
(3) Place both hands in the lukewarm water. Do both hands still sense the same temperature?
Conclusion: Your hand is a thermometric element, but it makes a poor thermometer!

Figure l0.3 (below)
As the liquid in the bulb expands or contracts because its temperature changes, the liquid level in the narrow tube rises or falls.

Figure l0.4 (below)   A bimetallic strip; the two metals (brass and iron) are tightly bound together with rivets and/or strong cement. Brass expands (or contracts) more when heated (or cooled) than does iron, and so the strip bends as shown.The longer metal is on the outside of the curve. Why is this?
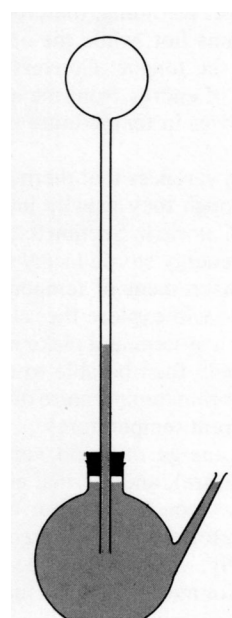

Figure 10.2 (above) Galileo's thermoscope. A glass bulb filled with air has a long tube whose end is dipped into water. When the bulb cools, the water rises in the tube; when the bulb is warmed, the water level drops.
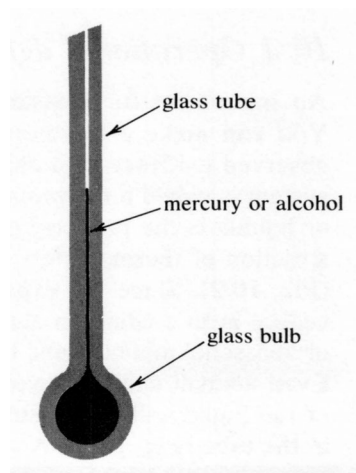

glass tube

mercury or alcohol

glass bulb

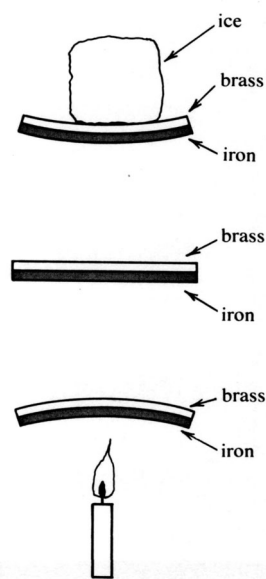ice
brass
iron

brass
iron

brass
iron

TABLE 10.1 A SCALE OF THE DEGREES OF HEAT*

Newton used two distinct operational definitions to define temperature (or "heat" as he called it). One made use of a linseed oil thermometer marked 0 at the freezing temperature of water and 12 at human body temperature. The other (used at higher temperatures, where the linseed oil thermometer failed) measured temperature by the time required for a particular hot block of iron to cool down to body temperature in a uniformly blowing cold wind. Newton conducted experiments that enabled him to convert from one scale to the other.

0	*The heat of the air in winter at which water begins to freeze. This heat is determined by placing a thermometer in packed snow while the snow is melting.*
0, 1, 2	*Heats of the air in winter,*
2, 3, 4	*Heats of the air in spring and in autumn.*
4, 5, 6	*Heats of the air in summer*
6	*Heat of the air in the middle of the day in the month of July.*
12	*The greatest heat that a thermometer takes up when in contact with the human body. This is about the heat of a bird hatching its eggs.*
17	*The greatest heat of a bath that can be endured for some time when the hand is dipped in it and is kept still.*
34	*The heat at which water boils violently...*
48	*The lowest heat at which a mixture of equal parts of tin and bismuth liquifies...*
96	*The lowest heat at which lead melts...*
136	*The heat at which glowing bodies shine at night but hardly in the twilight. . .*
161	*The heat at which glowing bodies in the twilight, just before the rising of the sun or after its setting, plainly shine, but in the clear light of day not at all or only very slightly.*
192	*The heat of coals in a little kitchen fire made from bituminous coal and excited by the use of bellows . . .*

*From Isaac Newton, *Opuscule,* 1701

(Fig. 10.3). A second common procedure is to make a so-called bi-metallic strip, which bends when its temperature is changed (Fig. 10.4). All thermometric elements have limitations, because the liquids solidify at very low temperatures and the solids melt at high temperatures.

 Temperature scales. Even though many temperature scales have been used in the past (see Table 10.1 for an example), and the Fahrenheit scale is commonly used in the United States today, we will describe the Celsius temperature scale, which is used for all scientific purposes. As we explained in Section 4.6, the temperatures of 0° and 100° on this

scale are equal, respectively, to the temperatures of a water-ice system (liquid water and ice at equilibrium) and a boiling water system (liquid water and steam at equilibrium) at ordinary atmospheric pressure. The question that remains is how to subdivide the interval between these two temperatures, and how to extend the scale above and below these temperatures. Each procedure leads to an operational definition of temperature.

Perhaps the simplest approach is to make a liquid-in-tube thermometer (Fig. 10.3), mark the stem for 0° and 100°, and then divide the interval between these marks into 100 equal spaces, each being one degree. Mercury, alcohol, solvents, and linseed oil have been employed as the liquid in thermometers. Since the scale constructed in this way depends on which liquid is used, one liquid has to be chosen as standard.

There are many other procedures that could be used. For example, a bimetallic strip (Fig. 10.4) could be used to operate a pointer that sweeps over a dial (Fig. 10.5). Marks are made on the dial for 0° and 100° as described above; then the interval is divided into 100 equal spaces. Or an energy-based scale could be constructed (Fig. 10.6).

The Celsius scale. The presently accepted scientific temperature scale has a formal definition remote from the simple operations described above. It evolved from the simple operational definition as the theories regarding temperature and energy transfer were developed during the nineteenth century.

We will be content with an operational definition (in box to left) making use of the mercury thermometer. This scale is very close to the energy-based scale (Fig. 10.5) as well as to the scientist's scale, which differ from one another by less than 1% in the magnitude of a degree interval. By contrast, the degree intervals on a Stoddard's solvent thermometer, which has been marked to read the Celsius scale defined above, are significantly non-uniform, being much larger near 100° Celsius than near 0° Celsius (Fig. 10.7).

The reference to the water-ice and water-steam equilibrium temperatures in the operational definition enables anyone to check the performance of a thermometer by testing it under these conditions. This is good practice because inexpensive commercial thermometers are not always reliable.

OPERATIONAL DEFINITION
Temperature is measured by the mercury level in a mercury thermometer, marked to read 0° in equilibruim with a water-ice system and 100° in equilibrium with a boiling water system (also called a water-steam system). The water-ice and water-steam systems must be at atmospheric pressure.

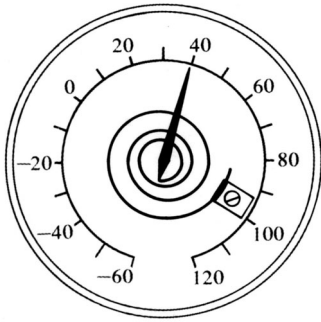*The symbol of temperature is T.*

Heat. When we discussed energy degradation in Section 4.1, the spontaneous tendency of systems at different temperatures to come to equilibrium at an intermediate temperature was mentioned. As these systems come to equilibrium, there is energy transfer from the hotter ones (energy sources) to the colder ones (energy receivers). As long as the systems interact and there is a temperature difference, there is energy transfer. Energy transferred because of a temperature difference has been given the special name *heat*. Energy sources, such as furnaces and candles, that function because of temperature differences, are called *heat sources*. For example, a lit match is a heat source for

Figure 10.5 A dial thermometer using a coiled bimetallic strip as a temperature-sensitive element.

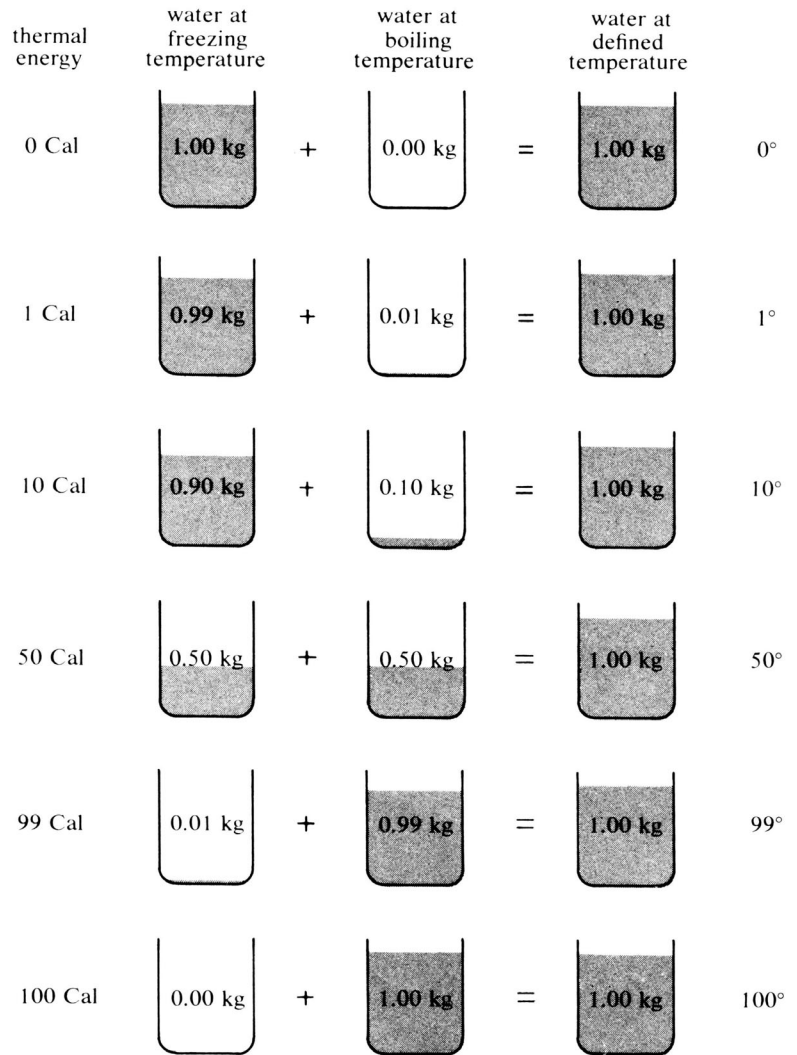| thermal energy	water at freezing temperature		water at boiling temperature		water at defined temperature	
0 Cal	1.00 kg	+	0.00 kg	=	1.00 kg	0°
1 Cal	0.99 kg	+	0.01 kg	=	1.00 kg	1°
10 Cal	0.90 kg	+	0.10 kg	=	1.00 kg	10°
50 Cal	0.50 kg	+	0.50 kg	=	1.00 kg	50°
99 Cal	0.01 kg	+	0.99 kg	=	1.00 kg	99°
100 Cal	0.00 kg	+	1.00 kg	=	1.00 kg	100°

Figure 10.6 Construction of an energy-based temperature scale. On this scale, the temperature (in degrees) of 1 kilogram of water is equal to its thermal energy in Calories.

Figure 10.7 The Celsius scale of a thermometer containing Stoddard's solvent.
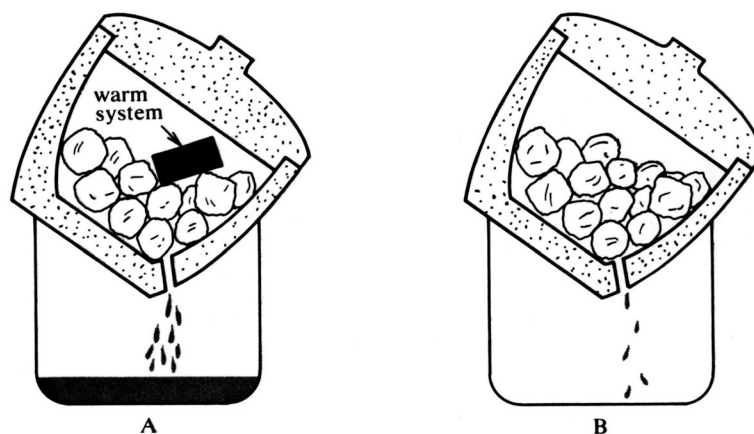
Figure 10.8 Apparatus for measuring the thermal energy of a warm or hot system. Insulated containers A and B are filled with ice at 0° Celsius. A warm system interacting with the ice in A melts some of it, in proportion to the system's thermal energy. Container B acts as a control; the amount of water melted in B is due to heat from the environment and must be subtracted from the amount melted in A.

an unlit candle, a glowing toaster coil is a heat source for the toast, human beings are heat sources in a cold room in the winter, and hot tea is a heat source for the ice cubes that convert it to iced tea. The functioning of a heat source depends on the temperature of the objects with which it interacts. Thus, the hot tea is a heat source with respect to ice cubes, but it is not a heat source with respect to the still hotter burner on the kitchen stove. For obvious reasons, heat input and heat output have an important effect on the temperature of systems.

10.2 Thermal energy and specific heat

You are now in a position to relate the thermal energy stored in a system to its temperature. According to the operational definition of temperature, you can find the temperature of the system with a Celsius thermometer. According to the operational definition of energy, you can find the energy of the system (in Calories) by letting it come to equilibrium with a water-ice system and weighing the additional melted water or solidified ice that is produced.

Measurement of thermal energy. The apparatus we will use for measuring the thermal energy of a system above 0° Celsius is shown in Fig. 10.8. It consists of two large, insulated containers of ice, each with a hole in the bottom through which water from melted ice can flow out. No water is included at the beginning. The system to be investigated is placed in one container A, and the other container B is used as "control" to determine the ice melted by interaction with room air in the absence of the system. To find the water produced by interaction of the system with ice, you subtract the mass of water coming out of the control container B from the mass of water coming out of the container with the system. Enough ice must be placed in the containers so not all of it is melted in the experiment. The time allowed must be adequate to cool the system to 0° Celsius. Thereafter, the water coming out of

TABLE 10.2 THERMAL ENERGY OF 1 KG OF WATER AT VARIOUS TEMPERATURES

Water temperature (Celsius)	Ice melted (kg)	Ice melted: control (kg)	Ice melted: net (kg)	Thermal energy (Cal)
10°	0.131	0.005	0.126	10
50°	0.625	0.003	0.622	50
75°	0.942	0.004	0.938	75

both containers is due to the melting of ice by its interaction with warm air in the room.

Results of using this apparatus to find the thermal energy of 1 kilogram of water and 1 kilogram of aluminum are listed in Tables 10.2 and 10.3 and are shown in Figs. 10.9 and 10.10. Two aspects of the results stand out: first, the graph of the experimental results is very close to a straight line; second, the thermal energy of 1 kilogram of water is between four and five times as large as the thermal energy of 1 kilogram of aluminum at the same temperature.

Specific heat. The observed straight-line relationship between thermal energy and temperature suggests a simple mathematical model (Section 1.3). The thermal energy of 1 kilogram of material is equal to the temperature multiplied by a *constant* number characteristic of the material (Eq. 10.1). This constant number is called the *specific heat* of the material (symbol C). The specific heat is therefore the energy transferred when 1 kilogram of material is raised or lowered in temperature by 1° Celsius. To find the specific heat of water, we use the data in the bottom line of Table 10.2: $C_{water} = (75\ \text{Cal})/(75°) = 1\ \text{Cal/°C}$. To find the specific heat of aluminum, we use the data from the bottom line of Table 10.3: $C_{aluminum} = (21.7\ \text{Cal})/(100°) = 0.22\ \text{Cal/°C}$. To obtain a more reliable value for the specific heat, we could make a large number of measurements and average the results.

So far we have discussed the thermal energy of 1 kilogram of material. With the help of a thought experiment in which you imagine 2 kilograms of a material interacting with the ice in a container, you can easily conclude that 2 kilograms contain two times the energy of 1 kilogram at the same temperature. The same holds true for larger and smaller amounts. In other words, the thermal energy is proportional to

Equation 10.1

thermal energy of 1 kg $= E$
temperature $\qquad = T$
specific heat $\qquad = C$

$E = CT$

(for water
$\qquad C = 1.0\ Cal/deg/kg)$
(for aluminum
$\qquad C = 0.21\ Cal/deg/kg$

(For an explanation of this mathematical model, which is a "direct proportion," see Section 1.3 and the Appendix—Mathematical Background, Section A.2.)

TABLE 10.3 THERMAL ENERGY OF 1 KG OF ALUMINUM AT VARIOUS TEMPERATURES

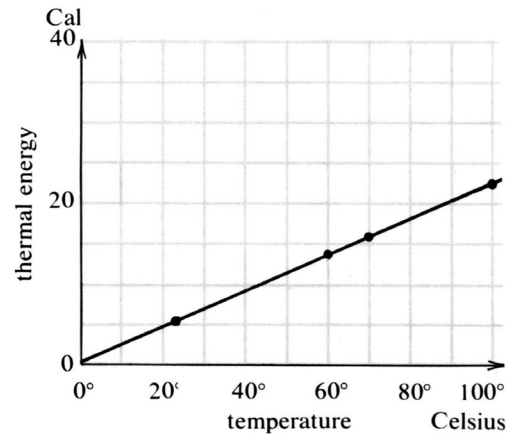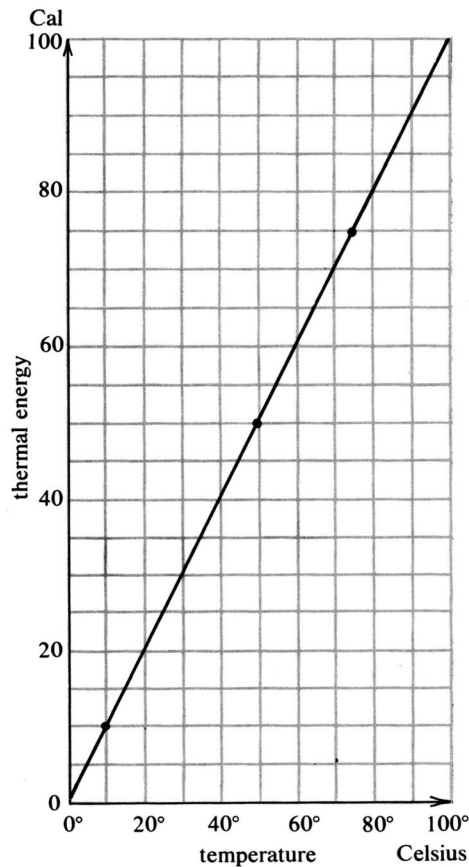| Aluminum temperature (Celsius)	Ice melted: net (kg)	Thermal energy (Cal)
24°	0.064	5.1
60°	0.163	13.0
75°	0.205	16.4
100°	0.271	21.7

Cal
100

Cal
40

Figure 10.10 (above) Thermal energy of 1 kilogram of aluminum at various temperatures.

Figure 10.9 (to left) Thermal energy of 1 kilogram of water at various temperatures.

Equation 10.2 (thermal energy)

thermal energy(Cal) $= E$
mass of material (kg) $= M_G$
specific heat (Cal/deg/kg) $= C$
temperature (°C) $= T$

$$E = CM_GT$$

the mass of the material. The combined proportionality to temperature and to mass is shown in Eq. 10.2. This mathematical model can be used for water, aluminum, or any other material for which the specific heat is known. The model is useful because it does apply to most materials at ordinary temperatures. If the model does not apply, and at extremes of temperature it does not, then additional measurements have to be made to find the thermal energy. Table 10.4, which contains specific heats of common materials, may be used for calculations of thermal energy according to the model of Eq. 10.2.

Applications. Most common systems are not made of one material. A glass of hot tea, for instance, is composed of hot water, a little dissolved tea material, and glass. A good model for the thermal energy of such a system is obtained by adding the energies of the various parts. This is done for a hot tea and glass system with the additional simplification that the energy of the small amount of dissolved tea is neglected (Example 10.1).

Subsystems at different temperatures. Sometimes a composite system includes subsystems that are initially at different temperatures. A glass of water with ice cubes is one example of such a system. Hot coffee and cold cream (initially in separate containers and not interacting) are another (Example 10.2). Boiling hot water and cold spaghetti are a third.

TABLE 10.4 SPECIFIC HEATS OF COMMON SUBSTANCES

Substance	Temperature* (Celsius)	Specific heat (Cal/deg C/kg)
air (gas)	0-100°	0.18
aluminum (solid)	-200°	0.08
	-100°	0.17
	0-100°	0.21
	100-600°	0.25
brass (solid)	0-100°	0.092
copper (solid)	0-100°	0.093
glass (solid)	20-100°	0.20
granite (solid)	10-100°	0.19
hydrogen (gas)	0-1000°	2.50
(liquid)	-252°	0.23
(solid)	-260°	0.57
iron (solid)	0-100°	0.110
lead (solid)	0-100°	0.031
oxygen (gas)	0-1000°	0.16
(liquid)	-200°	0.39
(solid)	-222°	0.34
sugar (solid)	20°	0.27
water (gas)	100-500°	0.37
(liquid)	0-100°	1.00
(solid)	-150°	0.25
	-100°	0.33
	-60°	0.39
	-20°	0.48
	-10-0°	0.50
wood	?	0.42
Wood's alloy (solid)	5-69°	0.035 to 0.05 (varies**)
(liquid)	70-150°	0.10 (approx.)

*Specific heats apply near given single temp. or throughout range.
**Varies depending upon composition of alloy.

Though the subsystems are initially at different temperatures, interaction among them will result in energy transfer, with the hotter ones acting as energy sources and the colder ones acting as energy receivers. The state of the system will change until the equilibrium state is reached with all subsystems at the same temperature. The mathematical model in Eq. 10.2 can be used to solve various problems associated with such an equilibrium state. For instance, what is the temperature of the cream and coffee system after the two subsystems are mixed?

You can estimate this temperature by treating the cup of coffee and the cream as an isolated system (Section 4.6). Therefore, the energy of the initial state of the system (before mixing) is equal to the energy of the final state (after mixing). Example 10.2 shows how to find the initial energy in joules. After mixing, the system comes to equilibrium. The coffee and cream are still sharing the same amount of energy as before, but they are now at the *same temperature*. Example 10.3 shows how to use this idea to calculate the final temperature of the mixture.

EXAMPLE 10·1. Thermal energy of tea–glass system.
Thermal energy of tea (assuming tea is water):

$$\text{tea } M_G = 0.25 \text{ kg}$$

specific heat of tea $C = 1.0$ Cal/deg/kg
temperature of tea $T = 70°$ Celsius

Energy of tea $= CM_GT$

$$= 1.0 \frac{\text{Cal}}{\text{deg-kg}} \times 0.25 \text{ kg} \times T$$

$$= 0.25 \frac{\text{Cal}}{\text{deg}} T$$

Thermal energy of glass:

$$\text{glass } M_G = 0.10 \text{ kg}$$

specific heat of glass $C = 0.20$ Cal/deg/kg
temperature of glass $T = 70°$ Celsius

Energy of glass $= CM_GT$

$$= 0.20 \frac{\text{Cal}}{\text{deg-kg}} \times 0.10 \text{ kg} \times T$$

$$= 0.02 \frac{\text{Cal}}{\text{deg}} T$$

Total thermal energy of system (E_{total}):

$$E_{\text{total}} = 0.25\text{T} + 0.02\text{T}$$

$$= 0.27 \frac{\text{Cal}}{\text{deg}} T$$

$$= 0.27 \frac{\text{Cal}}{\text{deg}} \times 70 \text{ deg}$$

$$= 18.9 \text{ Cal}$$

EXAMPLE 10·2. Thermal energy of hot coffee–cold cream system.
Thermal energy of coffee (assuming coffee is water):

$$\text{coffee } M_G = 0.30 \text{ kg}$$

specific heat of coffee $C = 1.0$ Cal/deg/kg
temperature of coffee $T = 60°$ Celsius

Energy of coffee $= CM_GT$

$$= 1.0 \frac{\text{Cal}}{\text{deg-kg}} \times 0.30 \text{ kg} \times 60 \text{ deg}$$

$$= 18 \text{ Cal}$$

Thermal energy of cream:

$$\text{cream } M_G = 0.05$$

specific heat of cream $C = 1$ Cal/deg/kg
(assumed to be like that of water)
temperature of cream $T = 5°$ Celsius

Energy of cream $= CM_GT$

$$= 1\,\frac{\text{Cal}}{\text{deg-kg}} \times 0.05 \text{ kg} \times 5 \text{ deg}$$

$$= 0.25 \text{ Cal}$$

Thermal energy of system (E_{total}):

$$E_{\text{total}} = 18 \text{ Cal} + 0.25 \text{ Cal} = 18.25 \text{ Cal}$$

EXAMPLE 10·3. Thermal equilibrium of cream and coffee.
Thermal energy of coffee:

$$\text{coffee } M_G = 0.30 \text{ kg}$$

specific heat of coffee $C = 1$ Cal/deg/kg
coffee temperature $T = $ (to be found)

Energy of coffee $= CM_GT$

$$= 1\,\frac{\text{Cal}}{\text{deg-kg}} \times 0.30 \text{ kg} \times T \text{ deg}$$

$$= 0.30 \; T \text{ Cal}$$

Thermal energy of cream:

$$\text{cream } M_G = 0.05 \text{ kg}$$

specific heat of cream $C = 1.0$ Cal/deg/kg
cream temperature $T = $ (to be found)

Energy of cream $= CM_GT$

$$= 1.0\,\frac{\text{Cal}}{\text{deg-kg}} \times 0.05 \text{ kg} \times T \text{ deg}$$

$$= 0.05T \text{ Cal}$$

Thermal energy of system (E_{total}):

$$E_{\text{total}} = 0.30T + 0.05T = 0.35T \text{ Cal}$$

From Example 10·2 you know that $E_{\text{total}} \approx 18$ Cal.
Therefore:

$$0.35T = 18 \text{ Cal}$$

$$T = \frac{18}{0.35}$$

$$T \approx 50°$$

Other uses of the model. The above model for thermal energy (Eq. 10.2, $E = CM_GT$) can be used in many other ways, too. One way is to find the specific heat of a material, such as the spaghetti dropped into hot water, from the cooling effect the spaghetti has on the water. Another is to find the rate at which the sun radiates energy to the surface of the earth from the heating effect of the sun on bodies of water. A third is to calculate how much cream must be added to coffee to achieve a desired temperature for the beverage. All of these examples illustrate how the mathematical model enables you to make many applications of data gathered in one standard situation, the determination of the specific heat, as summarized in Table 10.4.

10.3 Energy accompanying phase change

In Section 9.3 we used the melting of ice to define the thermal unit of energy. That is, the quantity of energy transferred to melt 0.0125 kilogram of ice was named 1 Calorie. This transfer of energy does not change the temperature of the water-ice system, it only changes the relative amounts of the solid and liquid phases of the material called "water" and thereby increases the phase energy of the water-ice system.

We introduced the term "phase energy" in Section 4.4 to denote the energy stored in a system by virtue of its phase. A gaseous material has more phase energy than the same material in liquid form, and a liquid material has more phase energy than the same material in solid form.

To learn more about the energy accompanying phase changes such as melting, solidification, evaporation, and liquefaction, we investigated the relation between energy and temperature of Wood's alloy, a mixture of bismuth, lead, tin, and cadmium. Wood's alloy has an especially low melting temperature for a metal, 69° Celsius. It is therefore easily possible to study its liquid phase and its solid phase. Figure 10.11 shows the data for an experiment in which the energy of 1 kilogram of Wood's alloy is measured at various temperatures. We describe how we made the measurements in the box at left.

Heat of fusion. The data points in Fig. 10.11 do not fit a single straight line (as in Figs. 10.9 and 10.10), but they do fit *two* straight lines, one for the temperature range (0-69°) where the alloy is solid,

We carried out the experiment with Wood's alloy in the same way as the experiments described above in Section 10.2 using water and aluminum. We heated the 1 kg metal sample to various temperatures and then allowed it to come to equilibrium with ice at 0°C. We measured the mass of ice melted and calculated the thermal energy transferred using Eq. 9.1 ($E = 80 M_G$). Finally, we plotted the results in Fig. 10.11.
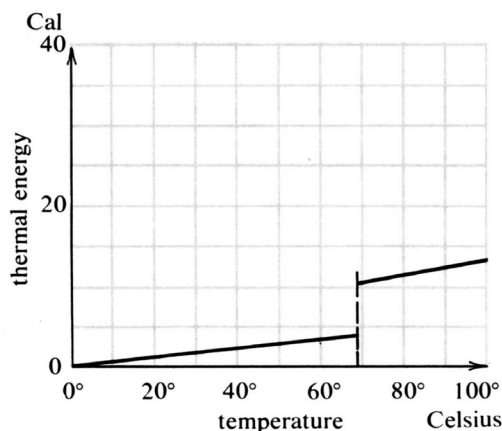

Figure 10.11 Apparent thermal energy of 1 kg Wood's alloy at various temperatures. The model of Eq. 10.1 does not fit the data points. The metal is solid below 69°C and liquid above that. At 69°C, you see a "step up" in the graph from 4 Cal to 11 Cal. Thus, the liquid phase at 69°C has 7 Cal more energy than the solid phase at the same temperature.

and the other for the temperatures (above 69°C) where it is liquid. At 69° Celsius there is a "step" in the graph. The significance of the step is that the liquid alloy at 69° has considerably more energy than the solid alloy at the same temperature. One kilogram of the liquid at 69° Celsius can melt 0.09 kilogram more ice than the solid alloy; this is simply another way of saying that it has about 7 Calories more energy than does the solid at the same temperature. The quantity of energy that is transferred when 1 kilogram of a material is melted or solidified without a temperature change is called the *heat of fusion.* The heat of fusion of Wood's alloy is 7 Calories per kilogram.

We therefore propose a mathematical model for the energy of Wood's alloy in which the alloy has thermal energy corresponding to a specific heat of about 0.05 Calorie per degree per kilogram in the solid phase (up to 69°C), an increase of phase energy equal to the heat of fusion of 7 Calories for each kilogram (at 69°C), and *additional* thermal energy corresponding to a specific heat of 0.10 Calorie per degree per kilogram in the liquid phase (which must be added if the sample is above 69°C). The algebraic form of this mathematical model is shown in the margin to the left. Calculations of the energy of Wood's alloy according to this model are in good agreement with the graph of the experimental results in Fig. 10.11 (Example 10.4).

Energy of ice systems and water systems. The heat of fusion of water is 80 Calories per kilogram by definition of the Calorie in Section 9.2.

EXAMPLE 10·4 The energy of 1 kilogram of Wood's alloy at 110° Celsius according to the model described in the text.
Data:

C (solid) = 0.05 Cal/deg/kg (up to 69°)
C (liquid = 0.10 Cal/deg/kg (above 69°)
heat of fusion = 7.0 Cal/kg (at 69°)

1 kilogram of solid at 40° Celsius:

$$E = CM_G T$$

$$= 0.05 \frac{\text{Cal}}{\text{deg-kg}} \times 1.0 \text{ kg} \times 40 \text{ deg}$$

$$= 2.0 \text{ Cal}$$

1 kilogram of solid at 69° Celsius:

$$E = CM_G T$$

$$= 0.05 \frac{\text{Cal}}{\text{deg-kg}} \times 1.0 \text{ kg} \times 69 \text{ deg}$$

$$= 3.5 \text{ Cal}$$

1 kilogram of liquid at 69° Celsius:

Modified form of Equation 10.2(including phase change, for energy of liquid Wood's alloy above melting temperature of 69°C).

(Eq. 10.2 in its original form, E = CMT, applies only *to situations where there is no phase change.)*

Temperature(deg C)   = T
Mass of liquid (kg)   = M
specific heat of solid
*   (Cal/deg C/kg)   = 0.05*
heat of fusion
*   (Cal/kg)   = 7.0*
specific heat of liquid
*   (Cal/deg C/kg)   = 0.10*

E = (0.05)M(69°C)
*   + (7.0)M*
*   + (0.10)M(T – 69)*

Comment:
The total energy (E) has three parts, each represented by a separate term in the above equation:

(0.05)M(69°C) is the thermal energy of the solid at 69°C.

(7.0)M is the phase energy of the liquid at 69°C.

(0.10)M(T – 69) is the additional *thermal energy gained by the liquid when it is heated above 69°C.*

E = energy of solid + heat of fusion

$= 3.5 + 7.0$

$= 10.5$ Cal

1 kilogram of liquid at 110° Celsius:

E = energy of liquid at 69° $+ CM_G(T - 69)$

$= 10.5 \text{ Cal} + \left[0.10\, \frac{\text{Cal}}{\text{deg-kg}} \times 1.0 \text{ kg} \times (110 - 69) \text{ deg} \right]$

$= 10.5 \text{ Cal} + \left[0.10\, \frac{\text{Cal}}{\text{deg-kg}} \times 1.0 \text{ kg} \times 41 \text{ deg} \right]$

$= 10.5 \text{ Cal} + 4.1 \text{ Cal}$

$= 14.6 \text{ Cal}$

". . . melting ice receives heat very fast, but the only effect of this heat is to change it into water, which is not in the least sensibly warmer than the ice was before. A thermometer, applied to the drops or small streams of water, immediately as it comes from the melting ice, will point to the same degree as when it is applied to the ice itself. . . A great quantity, therefore, of the heat, or of the matter of heat, which enters into the melting ice, produces no other effect but to give it fluidity . . . it appears to be absorbed and concealed within the water, so as not to be discoverable by the application of a thermometer."
  Joseph Black
  Lectures on the Elements
  of Chemistry, 1803

An energy of 80 Calories must be transferred to 1 kilogram of ice at 0° Celsius to melt it to water at 0° Celsius. Conversely, 80 Calories of energy must be transferred from 1 kilogram of liquid water at 0° Celsius to freeze it to 1 kilogram of ice at that temperature. It is clear, therefore, that 1 kilogram of ice and of water at 0° Celsius do not have the same energy from the viewpoint of energy conservation.

This conclusion is contrary to what is implied by the operational definition of energy in Section 9.2. Since no ice is melted while either ice or water at 0° Celsius comes to equilibrium with a mixture of water and ice, the energy of either system is 0 Calories according to our operational definition. The operational definition is therefore incompatible with energy conservation in its application to water and/or ice systems, and we must modify it. Therefore, we will now define the energy of liquid water at 0° Celsius to be 0 Calories. The energy of 1 kilogram of ice at 0° Celsius must then be -80 Calories (80 Calories less than zero). The energy of ice at temperatures below 0° Celsius is still less than this, as determined by the specific heat of ice (Table 10.4). The resulting model for the energy of water is similar to the model for Wood's alloy explained above and illustrated in Fig. 10.11. Can you sketch the graph or find an equation that gives the energy of water from below 0°C to above 100°C?

Heat of solution. Another important type of phase change, in addition to melting and vaporization, is the formation of a solution phase from a separate liquid phase and a solid phase, such as water and sodium chloride (table salt). This process is similar to the melting process in that the rigid solid material disappears and a new liquid phase is formed. There is a difference, however, in that the new liquid phase is a mixture of two substances rather than one pure substance. The energy of the solution phase is usually found (by application of the operational definition) to be different from the energy of the separate phases before mixing but at the same temperature. The energy difference is called the *heat of solution.* The heat of solution depends on both the amount of solid phase and the amount of liquid phase forming the solution. Since the solid undergoes the more drastic change, however, its amount has the more important influence; in the simplest mathematical model, the

heat of solution is directly proportional to the mass of the solid material dissolved and does not depend on the mass of liquid used.

Applications. Values of the energy transfer accompanying phase changes at the listed temperatures are given in Table 10.5. You can see that the energy required for melting sodium thiosulfate (photographer's "hypo") is closely equal to that required for dissolving it in water. This observation supports the point of view that melting and dissolving are similar processes.

In all the instances except the solution of sulfuric acid, the process tabulated is associated with an increase in the energy of the system. In other words, energy has to be supplied from some heat source (perhaps a burning candle, warm water, room air, or any other system at a higher temperature) in order to bring about the change at the listed tempera-ture. Energy has to be supplied to solid nitrogen at -210° Celsius to melt it. Energy has to be supplied to liquid alcohol at 78° Celsius to vaporize it. And energy has to be supplied to a system of solid sodium chloride and liquid water at 20° Celsius to produce a sodium chloride solution at 20° Celsius. Energy has to be removed from a sulfuric acid-water system at 20° Celsius when a solution is formed at the same tem-perature.

The reverse processes, of course, are accompanied by the opposite energy changes. In other words, liquid water freezing to form solid wa-ter at 0° Celsius acts as a heat source for colder systems (below 0° Cel-sius), gaseous water liquefying to liquid water at 100° Celsius acts as a heat source for colder systems (below 100° Celsius), and so on.

In all these examples, we have emphasized the temperature at which

TABLE 10.5 ENERGY ACCOMPANYING PHASE CHANGES

Substance	Process	Temperature (Celsius)	Energy increase (Cal/kg)
water	fusion	0°	80
water	vaporization	100°	540
mercury	fusion	-39°	2.82
mercury	vaporization	357°	65
nitrogen	fusion	-210°	6
nitrogen	vaporization	-195°	48
carbon dioxide	fusion	-56°	45
carbon dioxide	vaporization	-50°	85
alcohol	fusion	-114°	25
alcohol	vaporization	78°	204
sodium thiosulfate	fusion	46°	48
sodium thiosulfate	solution in water	20°	48
sodium chloride	solution in water	20°	22
sulfuric acid	solution in water	20°	-180

the phase change takes place. We have always compared two states of a system *at the same temperature* to avoid including thermal energy changes along with the phase energy changes. Table 10.5 lists the energy changes associated with a phase change (but no temperature change) of 1 kilogram of material.

Relation of phase change to temperature change. Frequently, phase changes and temperature changes occur together. For example, ice melts in a glass of warm water. The total energy of this system at the beginning of the process is approximately equal to the energy at equilibrium (some energy is exchanged with the room air), but the distribution of energy between the two subsystems has changed. Energy has been transferred from the warm water to the ice. An isolated system of 1 kilogram of ice at 0° Celsius in interaction with 1 kilogram of water at 80° Celsius will come to equilibrium in the form of 2 kilograms of water at 0° Celsius (Fig. 10.12). In this process heat flows from the warm water to the ice. The former cools down as energy is removed, the latter melts as energy is supplied.

The previous examples in which the melting of a solid substance (ice) was accompanied by a temperature drop in the isolated system certainly seemed natural to you. You may be a little surprised, however, when you mix water and sodium thiosulfate (both at room temperature, about 20° Celsius), shake the mixture to bring about phase equilibrium as the sodium thiosulfate dissolves, and then find that the temperature of the solution is only 10° Celsius (Fig. 10.13). How is it possible for two systems at the same temperature to come to equilibrium with one another

Figure 10.12 (to right) Ice at 0° Celsius and water at 80° Celsius interact to form water at 0° Celsius.
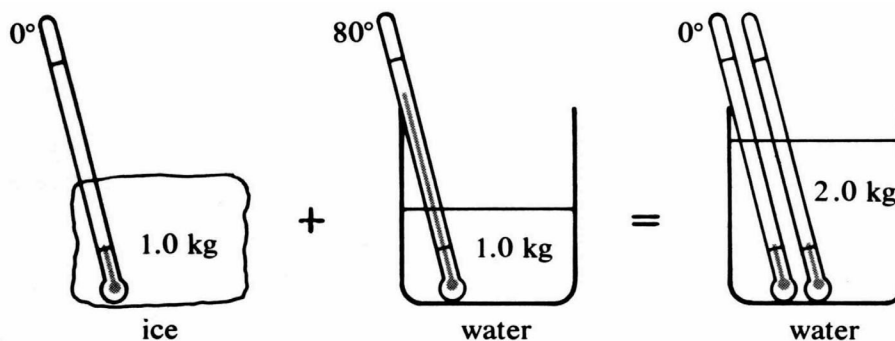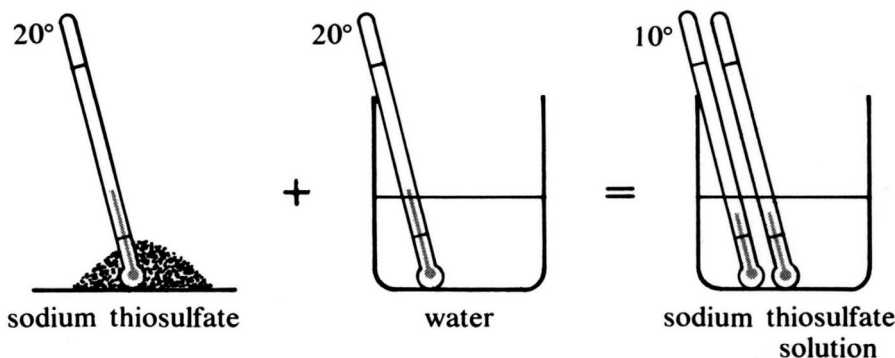
Figure 10.13 (to right) The temperature drop of the liquid is evidence that thermal energy of the system was transferred to phase energy while the sodium thiosulfate dissolved.
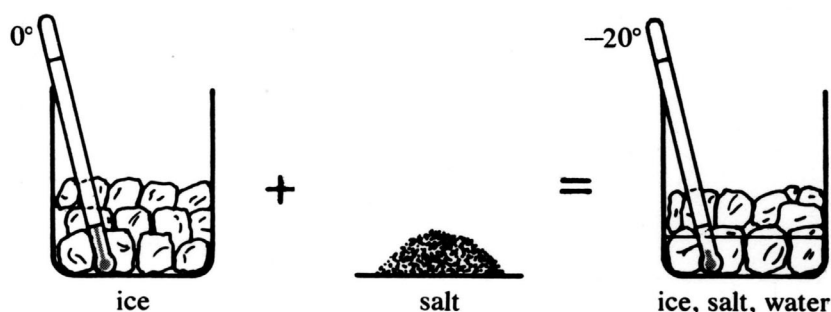
Figure 10.14 Salt (sodium chloride) sprinkled on ice interacts with the ice so that some of it melts. The temperature drop of the system is evidence that thermal energy was transferred to phase energy.

at a lower temperature? The answer is that the phase energy of the system increased as the sodium thiosulfate dissolved. Since the system in the jar was approximately isolated so that the total energy remained constant, most of the phase energy increase occurred at the expense of thermal energy stored in the system. Hence, the temperature decreased.

The same principle, transfer of thermal energy to phase energy, operates when ice cubes are dropped into cold salt water. The ice dissolves in salt water, its phase energy increases, and the thermal energy of the entire ice-water-salt system decreases, with a drop in temperature to as low as -20° Celsius. You can achieve even lower temperatures by sprinkling salt directly on ice cubes (Fig. 10.14). The salt dissolves in the water film on the surface of the ice to form a salt water phase, the ice then dissolves in the salt water phase with an increase in its phase energy, and the temperature of the entire system drops as its thermal energy is transferred to phase energy.

This phenomenon may startle you because you associate melting with an increase in temperature, freezing with a decrease. Here there was melting and decrease in temperature! The uncommon aspect of the experiments we have described was the *isolation* of the systems. Can you now explain why your commonsense expectation was misleading?

10.4 Chemical energy release

Chemical changes such as the burning of fuel are well known to result in the transfer of energy from the chemical form to the thermal form. The amount of energy can be measured by the melting of ice in a standard thermal system (water-ice mixture) and is found to be quite large, about 8000 Calories per kilogram for fuels such as gasoline, petroleum, candle wax, or coal.

Since burning fuel is an energy source, you can infer that the combustion products in the final state have a lower chemical energy than

TABLE 10.6 CHEMICAL ENERGY RELEASE

Fuel	Energy decrease of fuel-oxygen system (Cal/kg)
carbon	8,000
coal	5,000-8,000
wood	4,000
paraffin	10,000
alcohol	6,500
gasoline	11,500
kerosene	11,000
fuel oil	10,000
natural gas	11,000-13,000
hydrogen	30,000
iron	1,600
gunpowder	750
American cheese	4,500
egg	1,600
sirloin steak	2,500
rye bread	2,600
salad oil	9,000
sugar	4,000

did the fuel and oxygen in the initial state. The increase in temperature of the environment of the burning fuel is evidence of energy transfer. Table 10.6 lists the energy release accompanying combustion of various common fuels and foods.

One of the most significant technological achievements of the last 200 years is the transfer of chemical energy of fuel-oxygen systems to kinetic energy of steam engines, automobiles, airplanes, rockets, and other machinery. This process requires a suitably designed coupling element called a heat engine, which is described in Chapter 16.

10.5 Thermal energy and the MIP model for matter

The caloric theory. Greek philosophers, for whom fire was one of the elements of which all matter was composed, ascribed temperature differences to the presence of varying amounts of this element. Thus there originated the theory that heat (thermal energy) was a material substance, a "caloric" fluid.

This caloric theory, through the analogy of water flowing between reservoirs, enabled scientists to distinguish clearly between the temperature of a piece of material (represented by the level of the fluid) and the quantity of thermal energy (represented by the amount of fluid). The caloric fluid was endowed with properties not shared by water, air, or other physical fluids, in that the caloric could pass into and out of

matter. The caloric theory could explain thermal equilibrium (the caloric fluid moves from one body to another until all interacting bodies contain the same level), various specific heats for different materials (some could contain more caloric for the same change in level than others, as a wide bowl can hold more water than a narrow vase), and conservation of thermal energy (the caloric could move from one body to another, but it was neither created nor destroyed).

Refinements in the theory helped it to accommodate new information. For instance, the transfer of thermal energy to phase energy when ice melts without a temperature increase gave rise to the dual concepts of "sensible" and "latent" caloric. Latent caloric was so closely tied to the particles of matter in a body that it could not flow from one body to another and heat it. Sensible caloric could do so. On melting, sensible caloric was converted to latent caloric; on combustion, latent caloric was converted to sensible caloric.

One weakness of the caloric theory was that efforts to weigh the fluid were in vain. All experimental evidence indicated that changes in temperature and/or phase changes did not change the gravitational mass of a system. Caloric therefore became identified as a fluid that had no mass. This may be contrary to common sense, but it is of no real consequence in the definition of a working model as long as no other aspects of the model are contradicted.

Early forms of the kinetic theory. Some of the scientists in the seventeenth and eighteenth centuries did not accept the caloric theory. They took the view that thermal energy was a form of vibration or motion of the particles of material bodies. This theory is called the *kinetic theory*. In support of this theory, it was pointed out that thermal energy could be produced by motion, as through friction when you rub your hands together. Further support came from the related theory that gases exert pressure on their containers because they are made of rapidly moving particles that strike and recoil from the container walls.

The two theories, caloric and kinetic, competed until near the end of the eighteenth century. Then the experimental work of Joseph Black on the measurement of specific heats and other thermal properties of materials, which were so effectively explained by the caloric theory, caused that theory to be accepted almost universally.

Weaknesses of the caloric theory. We have already pointed out that the caloric fluid was thought to be mass less. Even though this was not a flaw in the logic of the caloric theory, it gave the fluid an aura of mystery not possessed by any ordinary material fluid. The amateur scientist Count Rumford therefore rejected the caloric theory.

A second strange quality convinced Count Rumford that the caloric theory was unacceptable: Rumford found that apparently unlimited amounts of caloric could be generated through friction. Rumford's observations were made during the boring of cannons, an activity he supervised at the military arsenal in Munich. He found that the cannon

Joseph Black (1728-1799), a doctor of medicine at the Universities of Glasgow and Edinburgh, was one of the discoverers of carbon dioxide. However, his important work in thermal physics, including measurements of specific heat and heat of fusion, was published posthumously under the title Lectures on the Elements of Chemistry *(1803). Black was an advocate of the caloric theory of heat.*

Count Rumford (1753-1814). This amazing American grew up as Benjamin Thompson in Rumford, Massachusetts (now Concord, New Hampshire). He left his native land after the Revolution because of his Tory sympathies. Ambitious and aggressive, he was knighted by King George III and was employed by the Elector of Bavaria as Chamberlain, Minister of War, and Minister of Police. At the Bavarian court, Thompson was made a Lord, and it is ironic and touching that the title he chose was the name of the town to which he could never return. Somehow Rumford found the time to initiate extensive experiments designed to clarify the nature of heat.

"It will perhaps appear to some of you somewhat strange that a body apparently quiescent should in reality be the seat of motions of great rapidity; but you will observe that the bodies themselves, considered as wholes, are not supposed to be in motion. The constituent particles, or atoms of the bodies, are supposed to be in motion, without producing a gross motion of the whole mass. These particles or atoms, being far too small to be seen even by the help of the most powerful microscopes, it is no wonder that we cannot observe their motion. . ."

*James Prescott Joule
Lecture at St. Ann's Church,
Manchester, England, 1847*

barrels and the metal chips separated from them became hotter than boiling water, and this continued as long as the boring continued. He also measured the specific heat of the metal chips and the cannon barrels and found it unaffected by the boring process, during which a great deal of latent caloric had allegedly been converted to sensible caloric. Rumford concluded that anything that could be produced in unlimited quantity without an apparent source could not possibly be a material substance.

Rumford's experiments, which were carried out at the beginning of the nineteenth century, appear to us to be very convincing evidence against the caloric theory. His contemporaries, however, thought otherwise. Rumford did not provide a new theory that could explain as many observations as effectively and plausibly as the caloric theory. Hence the caloric theory remained in vogue for another 50 years. It is curious that this period of time coincided with the period following Thomas Young's experiments (Section 7.2), before the wave model for light triumphed over the corpuscular model.

The modern theory of thermal and phase energies. Dalton's atomic theory for chemical reactions (Section 4.5) established the model that matter was composed of many interacting particles in the micro domain. Quantitative properties such as mass, size, and specific interactions made the atoms of the mid-nineteenth century quite different objects from the material particles about which scientists had speculated previously. It was therefore no longer so difficult to add the ideas that these particles had kinetic energy and that energy could also be stored in the fields by which they interacted-at-a-distance.

Thermal energy. James Joule connected the MIP model with thermal energy in the way still accepted today. He proposed that macro-domain thermal energy was stored as micro-domain kinetic energy and field energy of the interacting particles. The transfer of thermal energy from hotter to colder bodies (macro-domain view) was represented by the transfer of kinetic energy through collisions of the particles (micro-domain view). Joule therefore associated temperature with kinetic energy of the particles: the hotter the body, the greater the average kinetic energy of the particles; the colder the body, the smaller the average kinetic energy of the particles. When a hot and a cold body interact, the more energetic particles in the hot body transfer energy to the less energetic particles in the cold body, until the average kinetic energies (and therefore the macro-domain temperatures) are equal. The random motion of the individual particles cannot be detected in the macro domain because the particles are too small.

Phase energy. Joule ascribed macro-domain phase energy changes to changes in the micro-domain field energy. He concluded that the particles are rearranged as a body is melted, that the fields by which they interact are altered, and that more energy is stored in the fields in the liquid phase than in the solid phase. When thermal energy is converted

to phase energy (macro-domain view), kinetic energy is converted to field energy (micro-domain view). The converse occurs when a liquid material freezes to form a solid. Then phase energy is converted to thermal energy (macro-domain view), or field energy is converted to kinetic energy (micro-domain view).

Friction. The theory could also explain, in a way that Rumford had already anticipated, the increase of thermal energy through friction. During the relative motion of the hands rubbing against one another, the speeds of the particles and therefore their kinetic energies are increased. This topic is discussed more extensively in Section 11.8.

Gases. Some of the greatest quantitative successes of the new kinetic theory of thermal energy were achieved in its application to gases, for the following reason. We have already explained (Section 4.5) that gases in the MIP model are made of particles at such great distances from one another that they do not interact appreciably. Each particle is surrounded by the field it creates and is unaffected by the fields of the other particles. Consequently, the field energy stored in a gaseous system cannot change as long as the system remains a gas. All the thermal energy increases or decreases therefore coincide with increases or decreases of particle kinetic energy. This fact greatly simplifies the description of gases compared to solid and liquid materials, where complicated interactions do occur among the particles. In Chapter 16 we will describe a mathematical model that relates the temperature, pressure, thermal energy, and specific heat of gases to the properties of the particles in the MIP model. This achievement led to the universal acceptance of the kinetic theory and the abandonment of the caloric theory.

Summary

The most important message of this chapter is that thermal behavior of matter must be described by two variable factors: temperature and thermal energy. Temperature (in degrees Celsius) is defined operationally in terms of a mercury thermometer, which gives a visual indication by the expansion or contraction of the mercury in a slender glass tube. Temperature governs the direction of energy transfer, from hotter bodies, which act as heat sources, to colder bodies, which serve as energy receivers. If three bodies at different temperatures interact with one another, the one at the intermediate temperature receives energy from the hotter one and supplies energy to the colder one simultaneously. Two or more bodies at the same temperature are in thermal equilibrium, and no energy is transferred among them. Thermal energy may be transferred by physical contact of the interacting objects (when an ice cube is placed in hot water) or by radiation as an intermediary in interaction-at-a-distance (when the sun radiates to the earth).

Thermal energy of a system is defined operationally in terms of the

mass of ice *melted* when the system interacts and comes to equilibrium with a mixture of water and ice. However, if the system was sufficiently cold that water is *frozen* (not melted) during this process, then we must assign a *negative number* to the energy of the system. The unit of energy is the Calorie, the energy required to melt 0.0125 kilogram of ice.

The thermal energy of a system and its temperature are related: the higher the temperature, the greater the thermal energy. For many systems made of one material and with all parts at the same temperature, a simple mathematical model describes the relation adequately (Eq. 10.2). The specific heat, introduced as part of this model, is the thermal energy transferred when 1 kilogram of the material is raised or lowered 1° Celsius in temperature. Phase energy or chemical energy of a system can also be measured by the system's ability to melt ice.

The modern kinetic theory of thermal energy is based on the MIP model for matter. What we have defined as thermal energy in the macro domain is, according to the theory, stored as the total kinetic energy and as some of the field energy of the interacting particles in the MIP model for the system. What we have defined as phase energy in the macro domain is stored as the remaining field energy of the particles in the MIP model. What we have defined as temperature in the macro domain is related to the average kinetic energy of the interacting particles in the MIP model.

Equation 10.2

thermal energy (Cal) $= E$
mass of material (kg) $= M_G$
specific heat (Cal/deg C/kg)
$\qquad\qquad\qquad = C$
temperature (deg C) $= T$

$$E = C\,M_G\,T$$

"We may conceive then, that the communication of heat to a body consists, in fact, in the communication of (kinetic energy) to its particles."
James Prescott Joule
Lecture at St. Ann's Church Reading Room,
Manchester, England
1847

Additional examples

EXAMPLE 10.5. What is the specific heat of aluminum calculated from the data point at 60° Celsius in Table 10.3?

Solution: M_G = 1.0 kg, T = 60° Celsius, E = 13.0 Cal, E = CM$_G$T or

$$C = \frac{E}{M_G T} = \frac{13\,\text{Cal}}{1\,\text{kg} \times 60\,\text{deg}}$$
$$\approx 0.22\ \text{Cal/deg/kg}$$

EXAMPLE 10.6. What is the specific heat of liquid Wood's alloy from the following data: 1.0 kilogram of hot aluminum at 200° Celsius interacts with 2.0 kilograms of liquid Wood's alloy at 81° Celsius until the two metals come to equilibrium at a temperature of 144° Celsius?

Solution: Aluminum transfers energy as it cools from 200° to 144° Celsius (ΔT = 56° Celsius). The energy transfer ΔE and temperature drop ΔT are related by the mathematical model $\Delta E = CM_G\Delta T$.

For aluminum: $\Delta E = CM_G\Delta T = 0.22$ Cal/deg/kg x 1 kg x 56 deg
$\approx$ 12.5 Cal

This energy is transferred to the Wood's alloy, which is warmed from 81° to 144° (ΔT = 63° Celsius).

For Wood's alloy: $\Delta E = CM_G\Delta T$ or

$$C = \frac{\Delta E}{M_G\Delta T} = \frac{12.5\,\text{Cal}}{2\,\text{kg}\times 63\,\text{deg}}$$

$$= \frac{12.5}{126} \approx 0.10\,\text{Cal/deg/kg}$$

EXAMPLE 10.7. One kilogram of copper at 80° Celsius is placed in 0.10 kilogram of water at 10° Celsius. What is the equilibrium temperature of the isolated copper and water system?

Solution 1 (Algebraic method.) The total initial energy of the system is the copper energy plus the water energy.

copper energy: $E = CM_GT$
               = 0.093 Cal/deg/kg x 1.0 kg x 80 deg = 7.4 Cal

water energy: $E = CM_GT$
               = 1.0 Cal/deg/kg x 0.10 kg x 10 deg = 1.0 Cal

total energy: E = 8.4 Cal

The total final energy of the system is due to copper and water at the same final temperature (unknown).

copper energy: $E = CM_GT$
               = 0.093 Cal/deg/kg x 1.0 kg x T deg
                                        = 0.093T Cal

water energy: $E = CM_GT$
               = 1.0 x 0.10 kg x T deg   = 0.10 T Cal

total energy: E = 0.193 T Cal

energy conservation: 0.193 T = 8.4

final temperature: T = (8.4) / (0.193) = 43.5° Celsius

Solution 2 (Trial and error method). Energy will be transferred from the copper to the water until these two subsystems have come to the same temperatures. Use model

$$T = \frac{E}{CM_G}$$

Guess 1. Transfer 3.0 Calories from copper to water.

copper energy: 7.4 - 3.0 = 4.4 Cal

copper temperature: $T = \dfrac{E}{CM_G}$

$$= \frac{4.4\,\text{Cal}}{0.093\,\text{Cal/deg/kg}\times 1.0\,\text{kg}} = 47^\circ\ \text{Celsius}$$

water energy: 1.0 + 3.0 = 4.0 Cal.

water temperature: $T = \dfrac{E}{CM_G}$

$$= \dfrac{4.0\ Cal}{1.0\ Cal/\deg/kg \times 0.10\,kg} = 40°\ \text{Celsius}$$

More energy transfer is needed, but not much more.

Guess 2. Transfer 3.5 Calories from copper to water.
copper energy: 7.4 - 3.5 = 3.9 Cal

copper temperature: $T = \dfrac{E}{CM_G}$

$$= \dfrac{3.9\ Cal}{0.093\ Cal/\deg/kg \times 1.0\,kg} = 42°\ \text{Celsius}$$

water energy: 1.0 + 3.5 = 4.5 Cal

water temperature: $T = \dfrac{E}{CM_G}$

$$= \dfrac{4.5\ Cal}{1.0\ Cal/\deg/kg \times 0.10\,kg} = 45°\ \text{Celsius}$$

The copper temperature came out *below* the water temperature, which is physically impossible, indicating that Guess 2 was too high and that a little less energy transfer is needed. Since the water temperature is slightly too high and the copper temperature is slightly too low, we can estimate that the final equilibrium temperature will be about halfway between these two results, which is about 43.5°C. This technique of solving a problem by judicious use of estimating, plus what is called "successive approximation," may seem suspect, or even "unscientific." However, this method is quite powerful and generally useful, as it provides an independent check of a result found by use of an exact mathematical model. Most important, this method of successive approximation is what "real" scientists often use, because it works in circumstances where exact mathematical models do not exist or where finding an exact result just takes too long. In fact, most real-world problems do not have exact mathematical solutions, so making estimates and using successive approximation is the only method available.

List of new terms

thermal energy	heat	heat of solution
temperature	specific heat	chemical energy
thermometer	phase energy	caloric theory
thermometric element	heat of fusion	kinetic theory
Celsius temperature		

List of symbols

T	temperature (degrees Celsius)	ΔE	energy transfer
ΔT	temperature change	M_G	gravitational mass
E	energy	C	specific heat

Problems

1. (a) Carry out the experiment described in Fig. 10.1 and report your observations.
   (b) Describe an experiment designed to find out by how many degrees Celsius the sense of your hands can be "fooled."
   (c) Carry out the experiment described in (b) and report your result.

2. Galileo's thermoscope (Fig. 10.2) is sensitive to effects of temperature *and* of atmospheric pressure. Explain why this is the case. This fact was not recognized in Galileo's time.

3. The glass of the thermometer tube (Fig. 10.3) expands and contracts when it is heated or cooled. Describe the effect this circumstance has on the temperature readings of the thermometer.

4. Suppose the thermometer tube (Fig. 10.3) has a non-uniform inside diameter. What effect will this have on the temperature measurements?

5. Explain the functioning of a bimetallic strip as a thermometric element (Fig. 10.4).

6. Consider the device illustrated in Fig. 10.15. Describe how it could be used for the operational definition of temperature. Explain any precautions that must be taken.
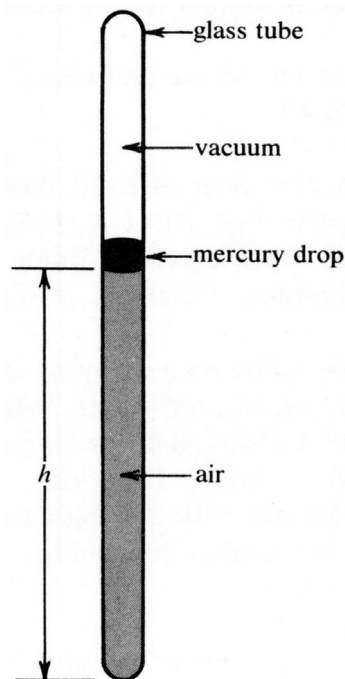

Figure 10.15 The sealed glass tube contains air on one side of a drop of mercury and a vacuum on the other. The mercury plugs the tube completely but can slide up and down.

7. Consider a thermometer constructed like Fig. 10.3, but using water as the thermometric element. Describe the limitations of this thermometer.

8. Find the thermal energy of each of these systems: (a) 1.5 kilograms of water at 66° Celsius; (b) 4.0 kilograms of wood at 45° Celsius; (c) 0.50 kilogram sugar at 52° Celsius; (d) 2.4 kilograms aluminum at negative 40 degrees Celsius (-40°C).

9. Find the temperature of each of these systems: (a) 1.8 kilograms of copper with a thermal energy of 11 Calories; (b) 6.0 kilograms of water with a thermal energy of 85 Calories; (c) 0.75 kilograms of wood with a thermal energy of -4.2 Calories; (d) 2.2 kilograms of Wood's Alloy with a thermal energy of 1.5 Calories. (Use values of specific heat from Table 10.4.)

10. Estimate the specific heat of the following materials from the data given.
    (a) 0.5 kilogram of ethyl alcohol initially at 60° Celsius melts 0.21 kilogram of ice at 0° Celsius.
    (b) 2.2 kilograms of hot stones at 300° Celsius melt 1.3 kilograms of ice at 0° Celsius.
    (c) 1.7 kilograms of paraffin at 45° Celsius melts 0.67 kilogram of ice at 0° Celsius.
    (d) When 1 kilogram of spaghetti at room temperature is dropped into 4 kilograms of boiling water, the water temperature drops to 90° Celsius.

11. Find the equilibrium temperatures of these systems:
    (a) 2.0 kilograms of water at 80° Celsius and 1.0 kilogram of water at 10° Celsius;
    (b) 1.0 kilogram of ice at 0° Celsius and 2.5 kilograms of water at 35° Celsius;
    (c) One dozen large eggs are taken from the refrigerator and put into 1 kilogram of boiling water. Assume that the flame under the boiling water is turned *off* immediately before you put the eggs into the water. Thus you can assume that the egg-water system is *isolated*. You will need to make and state your estimates for the temperature of the refrigerator and the mass of the eggs. *(Hint:* Treat the eggs as though they were made of water. To estimate the mass of the eggs, compare their apparent weight with something that you know, for example a quart of water or milk weighs about 1 kg. To estimate the temperature of the refrigerator, compare it with other temperatures that you know.)

12. Sprinkle salt on ice and report the lowest temperature you obtain (Fig. 10.14).

13. The specific heat of solid Wood's alloy found in the experiment reported in Fig. 10.11 is 0.050 Calorie per degree per kilogram, but that reported in Table 10.4 is 0.035 Calorie per degree per kilogram. Describe possible reasons for the discrepancy.

14. A family going on a camping trip was accustomed to taking along a gallon jug of cold water. When they filled it with 1 kilogram of ice at 0° Celsius and 3 kilograms of water at 10° Celsius, the ice lasted for 6 hours. How long should they expect the ice to last if they start out with 2 kilograms of ice and 2 kilograms of water? *(Hint:* Remember the energy transfer through the walls of the jug.)

15. Joseph Black, one of the physicists who contributed greatly to the concept of phase energy, recalled the following observation made by Fahrenheit in 1724: "He . . . exposed globes of water in frosty weather until . . . they were cooled down to the degree of the air, which was four or five degrees below the freezing point. The water . . . remained fluid so long as the glasses remained undisturbed, but on being taken up and shaken a little, a sudden freezing of a part of the water was instantly seen. . . . But the most remarkable fact is, that while this happens the mixture of ice and water suddenly becomes warmer, and makes a thermometer, immersed in it, rise to the freezing point." (Lectures on the Elements of Chemistry, Longman and Rees, London, 1803.) Explain Fahrenheit's observation in terms of phase energy.

Daniel Gabriel Fahrenheit (1686-1736) was the son of a German merchant. He was the first to introduce mercury as a thermometric substance and to succeed in making accurate thermometers. Fahrenheit is best remembered, of course, for the invention of the temperature scale that bears his name.

16. Explain how the following procedures or observations are related to phase energy.
(a) Salt is liberally mixed with the ice packed around an ice cream-making machine.
(b) A florist in Minneapolis wraps flowers to be delivered in very cold weather with soaking wet newspapers.
(c) A farmer puts a large tub of water into his cellar to keep the apples and potatoes stored there from freezing.
(d) You blow on hot soup to cool it more rapidly.
(e) Rubbing alcohol spilled on your hand feels cold.
(f) You feel cold when you step out of the swimming pool into a warm breeze.

17. Water is heated over a wood fire. How much tea can you make by burning 1 kilogram of wood under these conditions: you start with water at 10° Celsius; you want to heat the water to 80° Celsius only; only 10% of the energy from the fire is transferred to the water.

18. Warm-blooded animals maintain a steady state of body temperature generally above that of their environment. You can therefore represent a warm-blooded animal approximately as a passive coupling element, which converts chemical energy of a food-oxygen system to other forms of energy. Apply this concept to a Mr. X, who consumes food with an available energy of 2500 Calories per day.
(a) How much water evaporates from Mr. X's lungs and skin if all the energy becomes phase energy of gaseous water?
(b) How much hot air (warmed from room temperature of 20° Celsius to body temperature near 36° Celsius) can Mr. X produce in one day?

(c) How high must Mr. X climb if all the energy becomes gravitational field energy? His mass is 75 kilograms.

(d) Discuss qualitatively how the energy might be divided among these three and other possible forms. Comment whether you expect to find significant seasonal variations.

19. Which of the theories (caloric, kinetic) described in Section 10.5 seems most satisfactory to you for explaining the thermal phenomena described in Sections 10.1 to 10.4? Give reasons for your preference.

20. Interview four or more children (ages 8-14) to explore their concepts of thermal energy. Ask them to explain and/or predict their observations of some simple experiments, such as the mixing of two equal quantities of water of different temperatures or the melting of ice.

21. Interview four or more children (ages 8-14) to explore their concept of phase energy. Devise an appropriate demonstration and ask the children to explain and/or predict what happens.

22. Identify one or more explanations or discussions in this chapter that you find inadequate. Describe the general reasons for your judgment (conclusions contradict your ideas, steps in the reasoning have been omitted, words or phrases are meaningless, equations are hard to follow, . . .), and make your criticism as specific as you can.

Bibliography

S. C. Brown, *Count Rumford: Physicist Extraordinary,* Doubleday (Anchor), Garden City, New York, 1962.

S. G. Brush, *Kinetic Theory, Volume I: The Nature of Gases and Heat,* Pergamon, New York, 1965.

D. Cassidy, G. Holton, J. Rutherford, *Understanding Physics*, Springer Verlag, 2002.

C. A. Knight, *The Freezing of Supercooled Liquids,* Van Nostrand, Princeton, New Jersey, 1967.

E. Mach, *History and Root of the Principle of Conservation of Energy,* Open Court, La Salle, Illinois, 1911.

D. Roller, *The Early Development of the Concepts of Temperature and Heat: The Rise and Decline of the Caloric Theory,* Harvard University Press, Cambridge, Massachusetts, 1950.

J. Tyndall, *Heat, a Mode of Motion,* D. Appleton, New York, 1893.

A. Wood, *Joule and the Study of Energy,* G. Bell, London, 1925.

M. W. Zemansky, *Temperatures Very Low and Very High,* Van Nostrand, Princeton, New Jersey, 1964.

Articles from *Scientific American*

G. Y. Eastman, "The Heat Pipe" (May 1968).

W. Ehrenburg, "Maxwell's Demon" (November 1967).

D. M. Gates, "Heat Transfer in Plants" (December 1965).

D. Turnbull, "The Undercooling of Liquids" (January 1965).