4.4 Thermal Energy

When we were talking about mechanical energy, we were focused on doing work, that is, actually moving things around—force times distance and that sort of thing. Now we want to talk about another form of energy, which in everyday terminology we call *heat*. To be a bit more careful, we should call it *thermal energy*, but for our purposes simply calling it heat will do.

4.4.1 Temperature

We all have a pretty good idea of what temperature is, but defining it is a bit trickier. For starters, realize the adjectives we use to describe temperature are often pretty vague. We may say it is "hot" when it is 100°F outside, but if that is the temperature of our coffee we would probably say it was "warm." And if 100°F was the temperature of a charcoal briquette after a barbecue, we would describe it as being "cool." The context leads us to describe temperature with words, but just what does the numerical value mean? And just what is it that our measuring device (a thermometer, for example) is actually measuring?

Somewhere along the way, you have learned that the temperature scales measured with an ordinary thermometer are based on the temperatures of freezing water or, more precisely, an ice–water mixture, and the boiling point of water (at one atmosphere of pressure). On the **Celsius scale** (before 1948 it was called the centigrade scale), the freezing temperature is 0° and the boiling temperature is 100°. The Fahrenheit scale assigns the value of 32° to the freeze point and 212° to the boiling point. The relationship between the two is given by

Eq. 4.14
$$T(^\circ\text{F}) = 1.8\ T(^\circ\text{C}) + 32^\circ \quad \text{or} \quad T(^\circ\text{C}) = \frac{5}{9}\left[T(^\circ\text{F}) - 32^\circ\right]$$

The Fahrenheit and Celsius scales are based on extrapolating temperatures to any values no matter how high or how low. But in thermodynamics, there is merit in defining a temperature scale that isn't dependent on the properties of water and does have an absolute minimum value of zero, below which it is impossible to go. In classical physics, **absolute zero** corresponds to the point at which all molecular motion stops. However, quantum mechanically, molecules cannot cease all motion because that would violate the Heisenberg uncertainty principle, which asserts that you cannot ever know for sure what a particle is doing. So at absolute zero there will still be a certain small, but nonzero, energy known as the zero-point energy.

There are two measurement scales that use absolute zero as their reference temperature. In the SI system it is the Kelvin scale, named after Lord Kelvin (1824–1907). The temperature unit in this scale is the **kelvin** (not degrees Kelvin or °K; the degree designation was officially dropped in 1967). On this scale, absolute zero corresponds to –273.15°C (we'll just round that to –273°C). Note that the temperature interval on the Kelvin scale is the same as on the Celsius scale. That is, a temperature difference of 10°C is the same as a temperature difference of 10 K.

In the English system the temperature scale is named after William Rankine (1820–1872), and the units are designated as R (rankine). Absolute zero corresponds to –459.67°F (which we will round to –460°F), and each degree change on the Rankine and Fahrenheit scales is the same. That is, a change of 10 R is the same as a change of 10°F.

The following are handy conversions when using an absolute temperature scale:

Eq. 4.15 $$T(\text{K}) = T(°\text{C}) + 273 \quad \text{and} \quad T(\text{R}) = T(°\text{F}) + 460$$

4.4.2 Internal Energy, Thermal Capacitance

If we turn on the burner under a pot of water, we know that energy will be added to the water, raising its temperature. But there will be no work done, that is, no force times distance. That suggests that there are at least two ways to change the energy of a system: We can move something, doing work, or we can transfer heat, thereby changing the internal energy of the system.

Heat can be defined as the energy that is transferred between two systems (the stove and the pot of water) by virtue of a temperature difference. If we want to be picky, we shouldn't say an object has some heat in it, nor should we say we are adding some heat or taking some heat out of something. Objects don't contain heat, even after they have been heated up. What they do contain is **thermal energy**. Confusing? You bet. The technicality has to do with the *transfer* part of the definition; that is, heat exists only because of the temperature difference between two objects, in which case it is the energy that moves from one to the other. Fortunately, we don't want to be picky in this book, so you can relax and use your intuitive understanding of heat.

Potential and kinetic energy are observable, macroscopic forms of energy that are easily visualized and readily understood. More difficult to envision are microscopic forms related to the atomic and molecular structure of the system being studied. These microscopic forms of energy include the kinetic energies of molecules (that we measure with a thermometer) and the energies associated with the forces acting between molecules, between atoms within molecules, and within atoms. The sum of those microscopic forms of energy is called the system's **internal energy** and is represented by the symbol U. The **total energy** E that a substance possesses can be described then as the sum of its potential energy P.E., kinetic energy K.E., and internal energy U.

Eq. 4.16 $$E = U + \text{K.E.} + \text{P.E.}$$

The first law of thermodynamics can now be rephrased to say that for a **closed system**, that is, one in which we don't have to worry about matter passing through the boundary, if we add heat to the system (Q), and the system does work (W), the net result will be equal to the changes in the internal energy (ΔU), kinetic energy (ΔK.E.), and potential energy (ΔP.E.).

Eq. 4.17 $$Q - W = \Delta U + \Delta\text{K.E.} + \Delta\text{P.E.}$$

Quite often, it is changes in a substance's internal energy caused by changing its temperature that are of interest. For example, we may want to heat some water and store it in a tank so we can take a nice, hot shower, or we may want to design a house with lots of thermal mass to absorb incoming solar energy in the daytime, store it, and then give it back later to help keep the house warm overnight.

Table 4.3
Specific Heat and Volumetric Heat Capacity of Selected Substances

Substance	*Specific Heat*		*Density* (lb/ft^3)	*Heat Capacity* (Btu/ft^3 °F)
	(kJ/kg°C)	(Btu/lb°F)		
Water	4.18	1	62.4	62.4
Air (20°C)	1.01	0.24	0.081	0.019
Aluminum	0.90	0.22	168	37
Concrete*	0.88	0.21	144	30
Copper	0.39	0.09	555	50
Dry soil*	0.84	0.2	82	16
Gasoline	2.22	0.53	42	22
Steel*	0.46	0.11	487	54

* Representative values

For liquids and solids at atmospheric pressure, the change in energy stored (ΔE) when its mass (m) undergoes a temperature change (ΔT) is given by

Eq. 4.18 $$\text{Change in energy stored} = \Delta E = mc\Delta T$$

where c is called the **specific heat** of the substance. The specific heat is the energy needed to raise a unit of mass by 1 degree. For example, the specific heat of water is

Eq. 4.19 $$\text{Specific heat of water } c = 1 \text{ Btu/lb°F} = 4.18 \text{ kJ/kg°C}$$

Table 4.3 provides some examples of specific heat for several selected substances. It is worth noting that water has by far the highest specific heat of the substances listed; in fact, it is higher than almost all other common substances. This is one of water's very unusual properties and is in large part responsible for the major effect the oceans have on moderating temperature variations of coastal areas.

Also included in Table 4.3 are representative values of density along with the product of density and specific heat, which is known as the **volumetric heat capacity**. Volumetric heat capacity is an important concept in that it tells us how much thermal energy can be stored in a given volume of material as we raise its temperature. Quite often the design challenge is to find a way to pack away as much heat as possible in as small a space as you can. Notice again how water stands out as the substance that stores the most heat in the least volume—more than twice as much as concrete, which is the other most commonly used substance to store heat in passive solar houses.

For an example of how to use these concepts, see Solution Box 4.3.

4.4.3 Solids, Liquids, and Gases

The forces of attraction between atoms within a molecule are very strong. Molecules, in turn, exert small but appreciable attractions on one another, which are called **van der Waals attractions** (after nineteenth-century Dutch physicist Johannes van der Waals). That is, molecules are slightly "sticky." At low temperatures, molecules are rather docile; that is, they don't have much kinetic energy, and those sticky forces are sufficient to hold them pretty much in place in an orderly array. That is, the molecules are said to be in the **solid, crystalline state**.

When those molecules are heated up, they begin to vibrate more and at some point, called the **melting point** temperature, they have acquired sufficient energy to break loose from the crystal and can now begin to slide around past each other. These warmed molecules are still touching, but they are now mobile enough to fill out the shape of the container in which they are held. That is, the substance has made a transition from the solid state to the liquid state. The energy needed to melt a substance is called the **latent heat of fusion**. For example, the latent heat of fusion needed to melt ice is 333 kJ/kg (144 Btu/lb). It is also the amount of heat that has to be removed from the substance to cause it to go from the liquid to the solid state, that is, to cause it to freeze.

The ability of certain substances to absorb heat when they melt and give it back when they freeze makes possible the equivalent of a "thermal battery" for storing what is sometimes referred to as "coolth" (Figure 4.4). Making ice at night, when electricity rates are typically low, and then melting it the next day to provide air conditioning is one of the most cost-effective ways to cool buildings. (For example, see Solution Box 4.4.)

Ice is not the only phase-change material of interest in current efforts to reduce cooling loads in buildings. Materials that change phase at temperatures close to the desired indoor air temperature can be used to keep hot roofs from transferring daytime solar heat into interior spaces.

Having briefly considered the solid-to-liquid phase change, if still more heat is added, molecules can acquire so much kinetic energy that those puny van der Waals attractions can no longer compete, at which point the molecules can fly off in any direction they want to. This is the gas phase. A gas adapts to the shape of its container, just as liquids do, but gases differ in that they are easily expanded or compressed. As we shall see later when we talk about air conditioning systems, the expansion and compression of gases are of crucial importance. The temperature at which the transition from the liquid to gaseous state occurs is called the **boiling point** temperature, which for water is 100°C (212°F). The energy needed to do so is called the **latent heat of vaporization**. The latent heat of vaporization needed to cause water at 100°C to make the transition to steam at 100°C is 2257 kJ/kg (972 Btu/lb).

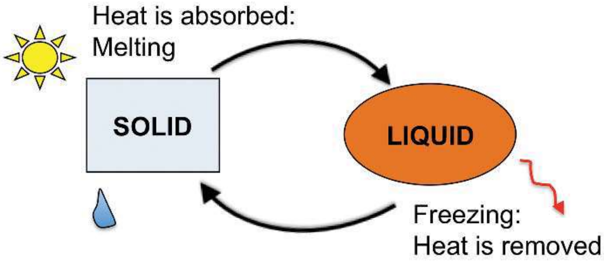

Figure 4.4 Phase-Change Materials Can Act Like Thermal Batteries