Chapter 3

Thermal Energy

In order to apply energy conservation to a falling ball or a roller coaster in the previous chapter, we had to assume that *friction* (with the air or the track) was negligible. Introduce some friction, and the falling ball's kinetic energy doesn't grow quite as quickly, while the roller coaster coasts to a stop, even on a level track. The effects of friction demonstrate that conservation of *mechanical* energy is not a universal law of physics.

But besides draining objects of their mechanical energy, friction also has another effect: It raises the *temperature* of objects. Everyone knows that you can warm up cold hands by briskly rubbing them together. Tires skidding along the road result in burnt rubber. A power drill's bit can quickly become too hot to touch. And although everyday falling objects don't often become noticeably warmer, the space shuttle requires state-of-the-art heat shields to prevent damage during its long, fast descent into the atmosphere.

Careful experiments show that the temperature increase of an object and its surroundings due to friction is directly related to the amount of mechanical energy lost. The most famous experimental demonstration of this fact was carried out by James Joule in the 1840's. Joule hung some weights from pulleys so that as they fell, they turned a paddle-wheel apparatus immersed in a bucket of water. The friction between the paddles and the water raised the water's temperature by an amount that was directly proportional to the distance that the weights fell. In our modern system of units, Joule found that raising the temperature of a kilogram of water by one degree Celsius required a loss in mechanical (gravitational) energy of approximately 4200 joules. Joule therefore proposed that this mechanical energy is not actually lost, but converted into a new type of energy: **thermal energy**, which manifests itself as an increase in temperature.

The temperature change caused by depositing a given amount of thermal energy in an object depends not only on the object's mass, but also on what it is made of. Raising the temperature of a kilogram of water by one degree Celsius requires 4200 joules, but for a kilogram of wood, only about 1700 joules are required. For iron the value is only 450 joules, while for lead, a mere 130 joules will suffice to accomplish the same one-degree temperature increase (per kilogram). Each of these numbers is called the **specific heat capacity** of the corresponding substance. In general, the specific heat capacity is the energy required to raise the temperature, per kilogram of material, per degree of temperature increase:

$$\text{Specific heat capacity} = \frac{\text{thermal energy input}}{(\text{mass}) \times (\text{temperature change})}. \tag{3.1}$$

To write this equation in symbols, I'll use C for specific heat capacity, T for temperature, and E_t for thermal energy. But the equation involves not T itself but the *change* in T during the energy-input process. The standard symbol for "change" is the Greek letter delta (Δ), so the change in T is written ΔT. Similarly, the thermal energy input is the amount by which the thermal energy changes, ΔE_t. Using these abbrevations, our equation becomes

$$C = \frac{\Delta E_t}{m \cdot \Delta T}. \tag{3.2}$$

Often it's useful to rearrange this equation to solve for the change in thermal energy:

$$\Delta E_t = m \cdot C \cdot \Delta T. \tag{3.3}$$

For example, to raise the temperature of a 10-kg wooden chair from 20°C to 25°C would require an energy input of

$$\Delta E_t = m \cdot C \cdot \Delta T = (10\text{ kg})(1700\text{ J/kg}\cdot{}^\circ\text{C})(5^\circ\text{C}) = 85{,}000\text{ J}. \tag{3.4}$$

(Note that the official units of C are joules per kilogram per degree Celsius.)

I should point out that the term "specific heat capacity" is in some ways a misnomer. It is most certainly not a *maximum* capacity in any sense: There is no limit on how much thermal energy a substance can absorb. Specific heat capacity is more analogous to the concept of capacity to hold one's liquor: a large "capacity" simply implies that it takes more liquor (or energy) to produce a noticeable change in behavior (or temperature). The use of the word "heat" in "specific heat capacity" is also somewhat misleading, as we'll see when we get to the technical definition of "heat" below. (The word "specific" simply means "per unit mass.")

Although there's nothing fundamentally special about water, it is essential to us humans. Moreover, our modern society uses quite a bit of energy to heat water, for various purposes. It's therefore a good idea to memorize the value of the specific heat capacity of water:

$$C_{\text{water}} = 4200\ \frac{\text{J}}{\text{kg}\cdot{}^\circ\text{C}} = 1\ \frac{\text{kcal}}{\text{kg}\cdot{}^\circ\text{C}} = 1\ \frac{\text{Btu}}{\text{lb}\cdot{}^\circ\text{F}}, \tag{3.5}$$

where the last two values follow from the definition of the kilocalorie and the definition of the Btu. Because quantities of water are often measured by volume instead of by mass, it's also useful to know that a kilogram of water occupies a volume of one liter (a little over a quart), while a pound of water occupies a volume of one pint (two cups, or half a quart). A gallon of water equals eight pints (pounds), or 3.8 liters (kilograms).

Exercise 3.1. Suppose you wish to make a nice hot cup of tea. Estimate the amount of energy required to heat the water, first in Btu's and then in kilowatt-hours. Use any reasonable values for the initial and final temperatures of the water.

Exercise 3.2. Suppose that you wish to have your hot cup of tea while camping in the wilderness, far from any convenient source of thermal energy, even firewood. No problem: You can simply put the cold water into a thermos bottle and shake the bottle vigorously, converting kinetic energy to thermal energy. Suppose that during each shake the water travels a distance of half a meter, and you shake it four times per second. What is the water's average speed during a shake? Roughly what is its maximum kinetic energy? Assuming that all this kinetic energy gets converted into thermal energy by the splashing water, how much will its temperature increase during one shake? How many shakes are needed to bring it up to boiling temperature? About how long must you continue shaking before enjoying your hot tea?

Exercise 3.3. A typical residential water heater holds 40 gallons. About how much energy (in Btu's) is required for it to heat a full tank of water? (You'll need to estimate the temperatures of your cold and hot tap water. If possible, use a thermometer to measure these temperatures.)

Exercise 3.4. Calculate the cost of the amount of energy computed in the previous problem, assuming that the energy comes from burning natural gas, which costs $7/MBtu (or use the actual cost from your utility bill).

Exercise 3.5. The human body contains a great deal of water, so for purposes of estimation, we can assume that its specific heat capacity is the same as that of water. Estimate the change in temperature that would result from putting one jelly donut (250 kcal) of thermal energy into your body. Comment on the result.

Exercise 3.6. The water in Malan's Falls drops a vertical distance of 300 feet. Consider just one kilogram of this water. How much gravitational energy does it lose during the fall? If all this energy gets converted into thermal energy of the water by the splashing at the bottom of the falls, by how much will the water's temperature increase? Comment on the result.

Exercise 3.7. Abel Knaebbel, professional stunt man and amateur physicist, dives from a 30-meter-high platform into a tub containing three cubic meters of water. Assuming that Abel's mass is 80 kg, how much gravitational energy does he lose during the dive? Now assume that all of this energy is converted into thermal energy by the splash, raising the temperature of the water. By how much does the water's temperature increase? (Hint: A cubic meter of water is 1000 liters.)

Phase Changes

When water boils, it continues to absorb energy (say, from the stove top), yet its temperature doesn't increase. Instead, the added energy causes the water to convert from a liquid into a gas (steam). This process is an exception to the rule that adding thermal energy raises an object's temperature. Any other **phase change**, in which a substance converts from one phase (solid, liquid, or gas) to another, is also an exception.

During a phase change, ΔT is zero but ΔE_t is nonzero, so the specific heat capacity ($\Delta E_t/m{\cdot}\Delta T$) is infinite (or undefined, as the mathematicians say). But the *total* amount of thermal energy required to cause a phase change, per kilogram of material, is still an important quantity. For instance, melting ice requires 80 kcal per kilogram, while boiling water requires 540 kcal per kilogram. Each of these numbers is called the **latent heat** of the corresponding phase change. In general,

$$\text{Latent heat} = \frac{\text{thermal energy input}}{\text{mass}} \qquad \text{(to accomplish the phase change)}, \quad (3.6)$$

or in symbols,

$$L = \frac{\Delta E_t}{m} \qquad \text{(to accomplish the phase change).} \qquad (3.7)$$

The latent heat values for the phase changes of water are, once again, 80 kcal/kg for melting ice and 540 kcal/kg for boiling water. To put these numbers into perspective, imagine starting with a kilogram of ice in a pan at 0°C, then gradually adding thermal energy to it. As you add the first 80 kcal, the ice melts into water, still at 0°C. Each of the next 100 kcal then raises the temperature of this water by one degree Celsius. At boiling temperature (100°C at sea level), the water starts to boil away into steam, but it takes another 540 kcal before the pan boils dry.

Another familiar transformation of water is *evaporation*, at or near room temperature. This process is essentially the same as boiling but less dramatic to watch: Again the end result is water in its gaseous phase. However, at room temperature the latent heat turns out to be somewhat greater than at boiling temperature: about 580 kcal/kg, instead of 540.

Latent heat values for the phase transformations of many other substances can be looked up in chemistry books, but are rarely needed in everyday life.

Exercise 3.8. Suppose that your stove top provides thermal energy to a pot of water at a rate of 2000 watts. After bringing the tea kettle to a boil and pouring yourself a cup, you put it back onto the stove, forgetting to turn the stove off. At this time the kettle contains half a liter of water. How long does it take for the kettle to boil dry? (Hint: First compute the energy required to boil away the remaining water, in joules.)

Exercise 3.9. Suppose that the temperature of your tap water is 15°C. You wish to cool a cup (250 g) of this water to 0°, so you add some ice. Assuming that the ice is already at 0°C, how many grams of ice must you add?

Exercise 3.10. Direct sunlight striking earth's surface provides energy at a rate of about 1000 watts per square meter of surface area. Suppose that the ground in a certain location (perhaps the mountains in springtime) is covered with snow to a depth of two meters, and that the snow is half ice and half air. Thus, each square meter of ground is covered with a cubic meter of ice—about 1000 kg. Assuming that the snow absorbs 10% of the energy in the sunlight, how many hours of direct sunlight are required to melt all the snow? Roughly how many weeks would this take, given that the sun is high in the sky for only a few hours per day?

Exercise 3.11. In Problem 2.4 you calculated the gravitational energy that you would gain upon climbing from the WSU campus to the summit of Mt. Ogden. However, because the human body is only about 25% efficient at converting chemical energy to mechanical energy, with the other 75% converted to thermal energy, you produce three JD of thermal energy for every JD of gravitational energy gained.

How much thermal energy is this? Assuming that all this thermal energy goes into evaporating water from your skin, how much water should you drink during the hike, to replace the lost fluids?

Heat vs. Work

Perhaps the most interesting property of thermal energy is that it spontaneously flows from one object to another, provided that the first object is hotter than the second. That's why ice melts in a warm room, fingers freeze on cold metal, and sunlight continuously warms the cool earth.

When energy spontaneously flows from a hotter object to a cooler object, we refer to that energy (while in transit) as **heat**. Heat is to be contrasted with **work**, which is any other type of energy transfer (such as exerting a force through a distance, as discussed in the previous chapter).

This use of the word "heat" takes quite a bit of getting used to. For instance, we should never say that a cup of boiling water "contains lots of heat," because once the energy is *in* the water, it doesn't matter whether the energy entered as heat or as work. If you put the water over a stovetop, then energy enters as heat. But you could vigorously shake or stir the water, with the same end result (hot water), accomplished this time by performing mechanical work. Or you could put the cup of water into a microwave oven: Since the flow of microwave energy from the magnetron source into the water is not caused by the magnetron being hotter than the water, we should classify this energy transfer as work (in this case, electrical work), not as heat.

Any transfer of energy from one object to another can be classified as either heat or work. This means that whenever the energy content of an object changes, that change must have been caused by heat flow, or work, or some combination of the two. For instance, if you add 2 MJ of heat to a pot of water and simultaneously perform 1 MJ of work, then the water's energy must increase by 3 MJ. In general,

$$\text{change in energy} = (\text{heat added}) + (\text{work done}), \tag{3.8}$$

or in terms of symbols,

$$\Delta E = H + W, \tag{3.9}$$

where E is the *total* energy content of an object, H is the amount of heat added to it, and W is the amount of work done on it (during some time period of interest). If heat or work is *extracted* from the object, then we simply use a negative value for H or W, and the same equation applies.

Exercise 3.12. Should the flow of electrical energy from the powerplant to your kitchen toaster be classified as heat or work? What about the flow of energy from the toaster's wires to your toast?

The Microscopic View of Thermal Energy

You now know how to calculate, or measure, the amount of thermal energy involved in various everyday processes such as boiling water. Still, you may be wondering, what *is* thermal energy? During the 1900's, many scientists asked this question, and eventually arrived at an answer.

Let's think some more about friction. Smooth surfaces, rubbed together, slide past each other with very little conversion of mechanical energy to thermal energy. Rough surfaces, on the other hand, exert a lot of frictional force on each other and convert mechanical energy into thermal energy much more rapidly. Why is this? Well, the bumps on the surfaces must be pushing against each other, then bending over, and finally springing back up as the move on. This pushing, bending, and springing converts the kinetic energy, from the overall motion, first into elastic energy and then partially back into kinetic energy of the tiny vibrations of the bumps. In other words, a uniform kinetic energy, with an object moving all in the same direction, is converted into kinetic and elastic energy that is irregularly distributed throughout the surfaces. This irregularly distributed energy gets passed back and forth among the particles that make up the surfaces, until its distribution is essentially random. This randomly arranged kinetic and elastic energy is what we call thermal energy.

The same sort of thing happens in Joule's paddle-wheel apparatus. The paddles stir the water, creating irregular whorls and vortices which rub against nearby bits of water, dispersing the energy into motions that are essentially random. Similarly, a parachute falling through the air sets up whorls and vortices that disperse energy into random air motions on a microscopic scale. Again, these random microscopic motions are what we call thermal energy.

All matter is made of tiny building blocks, called **atoms**. In a solid, the atoms are locked together in a regular geometrical structure; in a liquid or gas, the atoms are usually connected into small groups called **molecules**. In either case, the random microscopic motion and stretching that we call thermal energy is merely the ordinary kinetic energy of the atoms, and/or the elastic energy of atoms pushing against their neighbors.

It's interesting, though not really necessary, to ask how small these atoms and molecules are. The answer boggles the mind. The diameter of a typical atom is only a few tenths of a millionth of a millimeter, so it would take a few billion atoms, lined up in a row touching their neighbors, to span a full meter. A one-gallon jug of water measures about a billion atoms in each direction (width, depth, and height), so in total, it must contain roughly

$$(1,000,000,000) \times (1,000,000,000) \times (1,000,000,000) = 10^{27} \tag{3.10}$$

atoms. (Here 10^{27} simply means a 1 followed by 27 zeroes. Actually this estimate is a bit on the high side: the actual number of atoms in a gallon of water is only about 2/5 this much, or 4×10^{26}. Each water *molecule* happens to contain three atoms (two of hydrogen and one of oxygen), so the number of molecules in a gallon of water is "only" a little over 10^{26}.)

Please understand that I'm asking you to take my word for all this—I haven't offered you any evidence that atoms and molecules even exist, much less told you how to measure their sizes. Many of the best scientists doubted the existence of atoms until about 100 years ago, when evidence from a variety of difficult experiments all seemed to indicate that atoms truly do exist and have the approximate size that I've claimed. Unfortunately, the story of the discovery of atoms is one that we don't have time for now.

But if you're willing to take my word for all this, here's one more beautiful fact: In any gas or liquid at room temperature or higher, the average kinetic energy *per molecule*, due to random thermal motions from place to place, depends *only* on the temperature, not on what the material is made out of. (The same is true in solids, but with some ambiguity over what constitutes a molecule–an individual atom or a small group of them.) Furthermore, if we measure temperature on the absolute or **kelvin** (K) scale, so $T = 0$ means the lowest possible temperature ("absolute zero," or $-273°$C), then the formula that relates temperature to average kinetic energy is very simple:

$$\text{Average kinetic energy per atom} = \frac{3}{2}k_B T. \tag{3.11}$$

Here k_B is a univeral constant called **Boltzmann's constant**. As you might expect, the numerical value of k_B is tiny: 1.38×10^{-23} J/K. (10^{-23} means 1 divided by 10^{23}, or a decimal point followed by 22 zeroes and then a 1.) Individual molecules never have very much kinetic energy, even at very high temperatures. At room temperature, roughly 300 K, the average kinetic energy per molecule comes out to

$$\frac{3}{2}k_B T = \frac{3}{2}(1.38 \times 10^{-23}\ \text{J/K})(300\ \text{K}) = 6.2 \times 10^{-21}\ \text{J}.$$

Exercise 3.13. Calculate the total kinetic energy in a gallon of water, due to thermal motions of the water molecules. The approximate number of molecules in a gallon of water is given in the text.