ENERGY FOR SUSTAINABILITY

Technology, Planning, Policy

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CHAPTER 4

Fundamentals of Energy Science

4.1 Introduction

Before we can explore the array of technologies that will help us make the transition to more sustainable energy systems, we need a clear understanding of energy itself. Energy, primarily from the sun, flows through ecosystems and provides all living things, including our own bodies, the capacity to live, grow, repair tissue, reproduce, and do work. As it does so, its form may change from electromagnetic radiation flowing through space, to chemical energy stored in plants, to heat that keeps us warm, to potential energy as we climb a mountain, and to kinetic energy as we ski back down again. We use solar energy that plants captured and stored eons ago to heat our homes, generate our electricity, and drive our cars. As energy works its way through nature and human societies, it is constantly being transformed from one form to another. Although none is lost as it is stored, converted, and used, its quality is constantly being degraded to less and less useful forms, eventually ending up as relatively useless, low-temperature heat.

Understanding energy is one of the keys to understanding the universe and how physical and living systems work. In fact, a simple definition of *energy* is that it is the capacity to do *work.* Our own sensory perceptions and personal experiences have given us an inherent understanding of energy transformations and flows. We know about heat transfer (if only to put on a jacket or take it off), we know about the chemical energy released during combustion (as we accelerate to the next stoplight), we know about the energy of moving objects (if only to get out of the way of that oncoming car), and we know about radiant energy as we warm ourselves in front of that nice campfire. We even know something about the fairly esoteric concept of entropy, as our workshop gets more and more cluttered with sawdust and scraps as we build that nice piece of furniture for our home. Certainly, we have an intuitive understanding of many important energy systems, such as refrigerators, lights, cars, and furnaces, if only to operate them even though we may have only a vague notion of how they actually work.

Although a thorough explanation of the physics and chemistry of energy is well beyond the scope of this textbook, we can fairly easily develop an intuitive and somewhat quantitative feel for these energy transformations and flows. The vocabulary and basic principles presented here will provide the necessary foundation needed to understand the energy systems to be described in subsequent chapters.

The chapter begins with an introduction to the concept of energy itself, along with some units and conversions. It then explores some basic forms of energy, including mechanical, thermal, chemical, electrical, nuclear, and electromagnetic energy. What society cares about, of course, is not joules or British thermal units, but how we can transform various forms of energy into useful work to cool our beer, heat our homes, and take us where we want to go. The "rules of the road" that dictate what we can theoretically accomplish with a British thermal unit or a joule, as well as what we cannot do (such as create a perpetual motion machine) are introduced using the first and second laws of thermodynamics. With those fundamentals under our belts, we will be prepared for the following chapters, which explore some of the most important energy conversion systems.

4.2 Basics of Energy Science

Just what is energy? A precise answer to that deceptively simple question is surprisingly difficult. A common definition is that energy is "the capacity for doing work." Well, you and I are capable of doing work, does that mean we are energy? Although that may sound funny, Einstein's famous relationship between matter and energy, $E = mc^2$ says yes, we have mass, and mass and energy are inextricably linked to each other. But then, what is work? Work can be defined as the product of the force needed to move an object times the distance that it moves. But isn't thinking sometimes hard work? Moreover, work is not the only form of energy. For example, heat is another form of energy, but then what is heat? Well, heat is energy transferred from one object to another by virtue of their temperature difference. So, what is temperature?

Energy is a complicated concept. But, we can go far just by relying on our intuitive sense that energy is the ability to cause physical things to change. Energy allows us to make things get hotter, move faster, go uphill, and so forth.

4.2.1 Introduction to the First and Second Laws of Thermodynamics

Energy may change forms in any given process, as when chemical energy in wood is converted to heat and light in a campfire, or when the potential energy of water behind a dam is converted to mechanical energy as it spins a turbine, and then into electricity in the generator of a hydroelectric plant. The first law of thermodynamics says we should be able to account for every bit of energy in such processes, so that in the end we have just as much as we had in the beginning. With proper accounting, even nuclear reactions involving conversion of mass to energy can be accounted for.

To apply the first law, it is first necessary to define the system being studied. The system can be anything that we want to draw an imaginary boundary around—it can be a tree, or a nuclear power plant, or a volume of gas emitted from a smokestack. In the context of global climate change, the system could very well be the Earth itself. Quite often, what we really want to know is how efficiently a system converts energy in one form into useful energy in another form. For example, we might like to know how efficient a power plant is when it converts chemical energy in coal into electrical energy delivered to a transmission line. We can write the first law to express this as follows:

Eq. 4.1 **Energy into the system = Useful energy delivered + Wasted energy**

We want useful energy, which leads to the following definition of system efficiency

Eq. 4.2 $$\textbf{Energy efficiency } (\eta) = \frac{\textbf{Useful energy out}}{\textbf{Energy input}}$$

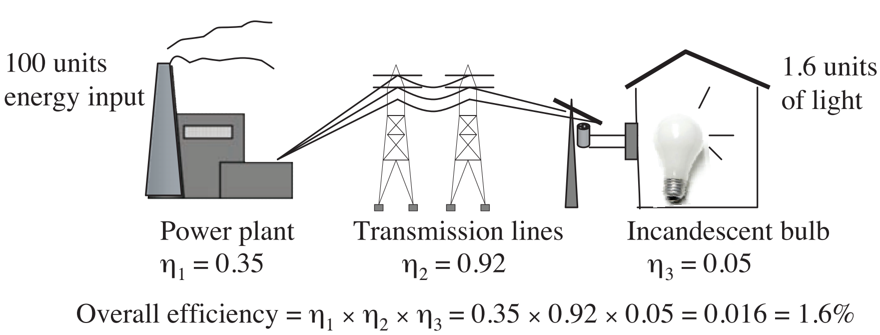
In practice, most of the systems we are interested in involve multiple energy transformations, each with its own efficiency. To find the overall efficiency from start to finish, we simply multiply the individual efficiencies. For example, consider a 35% efficient power plant putting electricity onto 92% efficient transmission lines that deliver electricity to a 5% efficient incandescent lightbulb. As Figure 4.1 indicates, that results in an overall efficiency of only 1.6%.

Figure 4.2 shows the benefits of switching those incandescent lights to energy-efficient, compact fluorescent lamps (CFLs). CFLs require only about one-fourth the electric power of an incandescent, while providing the same illumination. They also last as much as ten times as long. As shown, our 1.6 units of CFL light save a whopping 75 units of fuel into the power plant.

To summarize, then, the first law of thermodynamics tells us that energy is neither created nor destroyed as it makes its way through the universe. In other words, the first law gives us a bookkeeping system that allows us to keep track of *quantities* of energy.

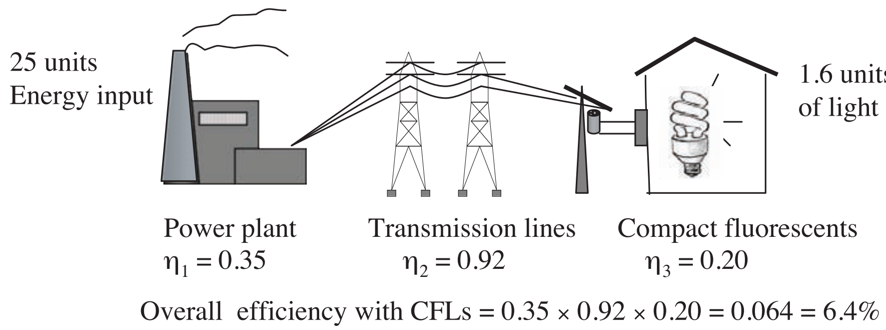
The second law of thermodynamics, on the other hand, tells us that even though no energy is "lost" during transformations, there will invariably be a loss in the *quality* of that energy. The quality of energy has to do with its capability to do useful things for us. For example, electricity is a very high-quality form of energy because it can do everything from powering your TV to heating your house. Low-temperature heat, on the other hand, is a very

figure 4.1



The overall efficiency from fuel to visible light from an incandescent bulb is a mere 1.6%.

figure 4.2



Switching from the incandescents in Figure 4.1 to energy-efficient compact fluorescent lamps (CFLs) saves three-fourths fuel needed at the power plant while delivering the same illumination.

low-quality form of energy. You may be able to warm your hands on that cup of coffee, but you certainly can't plug your computer into it. The second law tells us that no matter how hard we may try, every time we do something with energy there is always some loss in energy quality, which usually means some of it ends up as fairly useless waste heat.

We can imagine any number of processes that would satisfy the first law, but which we know realistically cannot occur. I can run some electricity through a heating element to heat my coffee, but I can't heat the element and expect to be able to get an equivalent amount of electricity back again. The second law takes care of such concerns by informing us about the direction in which processes can go. Electricity to heat is easy; heat to electricity is not.

The implications of the second law are profound. It "outlaws" a whole bunch of things, ranging from perpetual motion machines to the possibility of a warm cup of coffee heating itself up by stealing heat from the cool air in your kitchen. It dictates the maximum possible efficiency of your current automobile's engine as well as the fuel cell that may someday replace it. It tells us strange things about continuously increasing disorder in the universe. While you are cleaning up your room, making it more orderly, the power plant making electricity for your vacuum cleaner is creating even more disorder elsewhere as it converts an organized chunk of coal into disorderly gaseous and particulate emissions from its stack. We will explore the second law more carefully in Chapter 9, which discusses heat engines and power plants.

4.2.2 A Word about Units

William Thomson (Lord Kelvin) once stated that

> When you can measure what you are speaking about, and express it in numbers, you know something about it; but when you cannot measure it, when you cannot express it in

numbers, your knowledge is of a meager and unsatisfactory kind; it may be the beginning of knowledge, but you have scarcely, in your thoughts, advanced to the stage of science.

In the United States, energy units are often reported in both the English system (or, as it is sometimes called, the U.S. Customary System), and the International System (SI), so it is important to be familiar with both. In the SI system, the units of mass, length, and time are the kilogram (kg), meter (m), and second (s). In the English system, they are the pound-mass (lbm), foot (ft), and second (s).

Notice, we have already introduced some potential for confusion. What is this pound-mass business? Isn't a pound a pound? In common American usage, a *pound* is a unit of force (lbf), not a unit of mass. When we say something weighs 6 pounds, we are referring to the force it exerts on a scale, not its mass. If we took that 6-pound object to the moon, where the gravitational force is much lower, it would weigh only about 1 pound. Its mass would, however, be the same on Earth or the moon. As long as we stay at sea level on the Earth's surface, however, a pound is a pound. That is, a mass of 1 lbm does weigh 1 lbf.

The SI system avoids this confusion by always referring to the mass of an object, not its weight. Thus a 1 kg object on Earth has the same mass as a 1 kg object on the moon (see Figure 4.3).

Newton's second law connects mass (m) and weight (W) with the local gravitational acceleration, which is 9.807 m/s^2 or 32.174 ft/s^2 at the Earth's surface. For computational purposes, we'll call the acceleration g and just round it to 9.8 m/s^2 and 32.2 ft/s^2. Newton's equation is

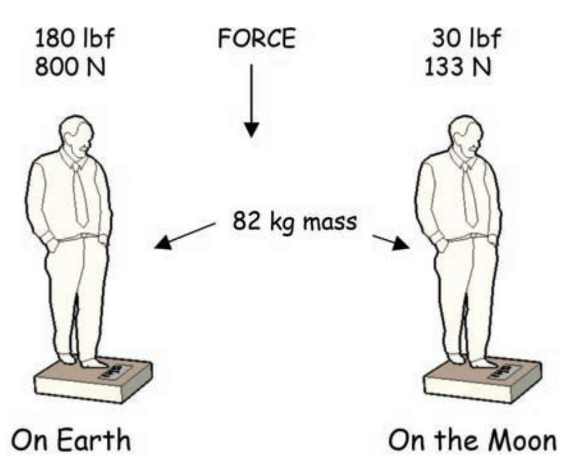
$$\text{Weight} = mg$$

Eq. 4.3

In SI units: $\text{Weight (Newton)} = \text{mass (kg)} \times g\ (\text{m/s}^2)$

In English units: $\text{Weight (lbf)} = \text{mass (slugs)} \times g\ (\text{ft/s}^2)$

figure 4.3



A man weighing 180 pounds on Earth will weigh only about 30 pounds on the moon. His mass does not change, however.

If you were to exert a force of 1 Newton on a scale on the Earth's surface, it would register 0.2248 lb, which is about the weight of one small apple. Curious coincidence, isn't it? Newton also gave us his amazingly important relationship between force, mass, and acceleration:

Eq. 4.4 **Force = mass × acceleration or F = ma**

$$\text{Force} = \text{mass} \times \text{acceleration} \quad \text{or} \quad F = ma$$

A force of 1 Newton will cause a mass of 1 kg to accelerate at the rate of 1 m/s^2.

Work, which is a form of energy, can be defined as force times distance. In SI units, force is in Newtons (N) and distance is in meters (m). The product, Newton-meters, is defined as joules:

1 joule = 1 Newton-meter

In English and American units, energy is quite likely to be measured in British thermal units (Btu), where a Btu is the energy required to raise 1 pound of water by one degree Fahrenheit. An older heat unit is the calorie, which is defined as the quantity of heat required to raise 1 gram of water by 1 degree Celsius. A calorie is about four times the size of a joule. Notice the interesting difference between the SI interpretation of energy as force times distance, whereas the English system based on the calorie suggests heat as a measure of energy.

4.2.3 The Distinction between Energy and Power

We all have pet peeves, and the constant misuse of energy and power units in popular media can be especially annoying. When we read about a new power plant described as delivering thus and such kilowatts per year, it is a sure tip-off that the writer is not very energy literate. Because you don't want to be considered illiterate, let's get the vocabulary right.

Power is energy per unit of time. It is a rate. For example, in SI units, power is often given in joules per second (J/s). One joule per second is designated as one watt in honor of the Scottish engineer James Watt, who developed the reciprocating steam engine.

1 watt = 1 J/s = 3.412 Btu/hr

Thus, to describe something in kilowatts per year is like some sort of energy acceleration unit, joules per second squared. It makes no sense at all. An electric heater that uses 10 kilowatts of power for two hours uses 20 kilowatt-hours (20 kWh) of energy. A water heater that uses natural gas at the rate of 16,000 Btu per hour (power) for one-half hour burns 8000 Btu of gas (energy).

Table 4.1 presents some unit conversion factors for both energy and for power. And because numbers can range from extremely small quantities (e.g., nanometers) to extremely large ones (e.g., exajoules), it is handy to have a system of prefixes to accompany the units. Some of the most important prefixes are presented in Table 4.2.

table 4.1 Energy and Power Units

ENERGY	1 British thermal unit	= 778 ft-lb
		= 252 calories
		= 1055 joules
		= 0.2930 watt-hours
	1 quadrillion Btu	= 10^{15} Btu
		= 1055×10^{15} J
		= 2.93×10^{11} kWh
		= 172×10^{6} barrels (42 gal) of oil equivalent
		= 36.0×10^{6} metric tons of coal equivalent
		= 0.93×10^{12} ft^3 of natural gas equivalent
	1 joule	= 1 Newton-meter (N-m)
		= 9.48×10^{-4} Btu
		= 0.73756 ft-lb
	1 kilowatt-hour	= 3600 kJ
		= 3412 Btu
		= 860 kcal
	1 kilocalorie	= 4.185 kJ
POWER	1 kilowatt	= 1000 J/s
		= 3412 Btu/hr
		= 1.340 hp
	1 horsepower	= 746 W
		= 550 ft-lb/s
	1 quadrillion Btu per year	= 0.471 million barrels of oil per day
		= 0.03345 terawatt (TW)

4.3 Mechanical Energy

Energy exists in many forms and a major task for engineers is to design systems that transform it from one form to another. We may want to use sunlight (electromagnetic energy) to run our TV (electrical energy), convert gasoline (chemical energy) into motion (kinetic energy), or cause uranium to fission (nuclear energy) to create steam (thermal energy) to run a turbine (rotational energy). Before we can explore such systems we need a brief introduction to these various forms of energy.

The energy systems that are most familiar to us from high school physics classes are often those based on moving weights around—pick them up to gain potential energy, drop them to demonstrate kinetic energy, rotate a wheel to demonstrate gyroscopic forces, and so forth. Those examples of mechanical energy are not only intuitively understandable, but they are also

table 4.2 Common Unit Prefixes

Quantity	Prefix	Symbol
10^{-12}	pico	p
10^{-9}	nano	n
10^{-6}	micro	μ
10^{-3}	milli	m
10^{-2}	centi	c
10^{-1}	deci	d
10^{3}	kilo	k
10^{6}	mega	M
10^{9}	giga	G
10^{12}	tera	T
10^{15}	peta*	P
10^{18}	exa	E

* In the United States, "quad" (short for quadrillion) is often used.

easy to analyze and provide an excellent introduction to the astonishing contributions that Sir Isaac Newton made to our understanding of physical science more than 300 years ago.

4.3.1 Potential Energy

It takes energy to lift a weight from one elevation to another, and in the process the object acquires potential energy; that is, it has the potential to do some work if we drop it. That energy required to raise an object, against the force of gravity, is described by the force times distance relationship. The force needed is the object's weight, which is the same as its mass (m) times the local gravitational acceleration (g). When raised to a height (h), its potential energy becomes

Eq. 4.5 $$\text{Potential energy (P.E.)} = \text{force} \times \text{distance} = \text{weight} \times \text{height} = mgh = Wh$$

Solution Box 4.1 shows how the units work in both SI and the English system.

4.3.2 Kinetic Energy

If we were to drop a mass (m) from a height (h) in a vacuum (so there is no air friction to slow it down), just before it hit the ground it would be moving at a speed of

Eq. 4.6 $$v = \sqrt{2gh}$$

SOLUTION BOX 4.1

Potential Energy Gained at the Top of the Mountain

How much potential energy is acquired when a 70 kg (154 lb) man reaches the top of a 2000 m (6562 ft) mountain? How many donuts of energy is that if 1 donut = 150 Calories? Note that food "Calories" (with a capital C) are actually kilocalories of energy.

Solution:

Using Equation 4.5 with SI units, we have

In SI: $\mathbf{P.E. = 70\ kg \times 9.8\ m/s^2 \times 2000\ m = 1.37 \times 10^6\ J}$

In English: $\mathbf{P.E. = 154\ lb \times 6562\ ft = 1.01 \times 10^6\ ft\text{-}lb}$

Using Table 4.1, we find that 1 kc = 4.185 kJ, so the donut equivalent is

$$\frac{1.37 \times 10^6\ \text{J}}{150\ \text{kc/donut} \times 4.185\ \text{J/kc}} = 2.2\ \text{donuts}$$

SOLUTION

At that point, all of the object's initial potential energy (P.E.) would be converted to kinetic energy (K.E.), so by substituting the value of h given in Equation 4.6 into Equation 4.5 we get the familiar equation for kinetic energy:

Eq. 4.7

$$\text{P.E.} = mgh = mg\left(\frac{v^2}{2g}\right) = \frac{1}{2}mv^2 = \text{K.E.}$$

4.3.3 Pressure Energy

The defining characteristic of energy is that it enables work to be done. A gas under pressure can force a piston to move and the pressure of water behind a dam can cause a turbine to spin, so pressure is, in essence, another form of potential energy.

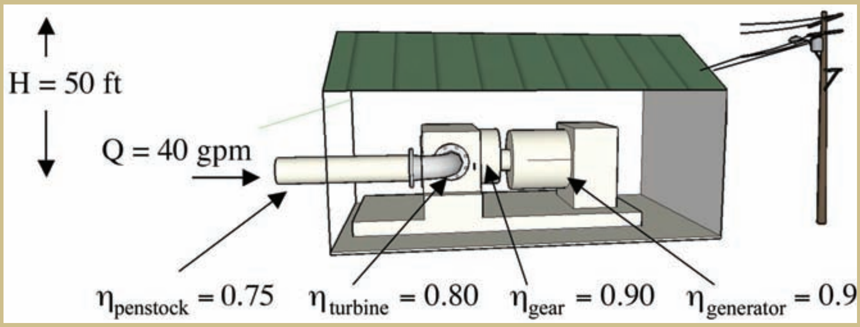
The hydroelectric system shown in Figure 4.4 illustrates all three forms of mechanical energy just introduced. The energy in a hydroelectric system starts out as potential energy by virtue of its height above the powerhouse. Water under pressure in the penstock (connecting pipe) is able to do work when released, so there is energy associated with that pressure as well. Finally, as water flows, there is the kinetic energy of moving mass. The hydroelectric system transforms energy through all three forms, from potential, to pressure, to kinetic energy. A turbine and generator in the powerhouse convert that energy to electrical power.

SOLUTION

SOLUTION BOX 4.2

Tapping into That Little Spring

Suppose you have a small spring up the hill from your cabin. You estimate it to be 50 feet higher than your house and you think you might be able to deliver about 40 gallons per minute to the powerhouse down near the cabin. You estimate the efficiencies of penstock, turbine, gearing, and generator to be 75%, 80%, 90%, and 90% respectively. How much power would your generator deliver and how much energy would it provide per month?



Solution:

From Equation 4.9, the overall efficiency of the hydropower system is

$$\eta = \eta_{penstock} \cdot \eta_{turbine} \cdot \eta_{gears} \cdot \eta_{generator} = 0.75 \times 0.80 \times 0.90 \times 0.90 = 0.49 = 49\%$$

and from Equation 4.8:

$$P(\mathrm{W}) = \frac{\eta Q(\mathrm{gpm})\, H(\mathrm{ft})}{5.3} = \frac{0.49 \cdot 40 \cdot 50}{5.3} = 185\ \mathrm{W} = 0.185\ \mathrm{kW}$$

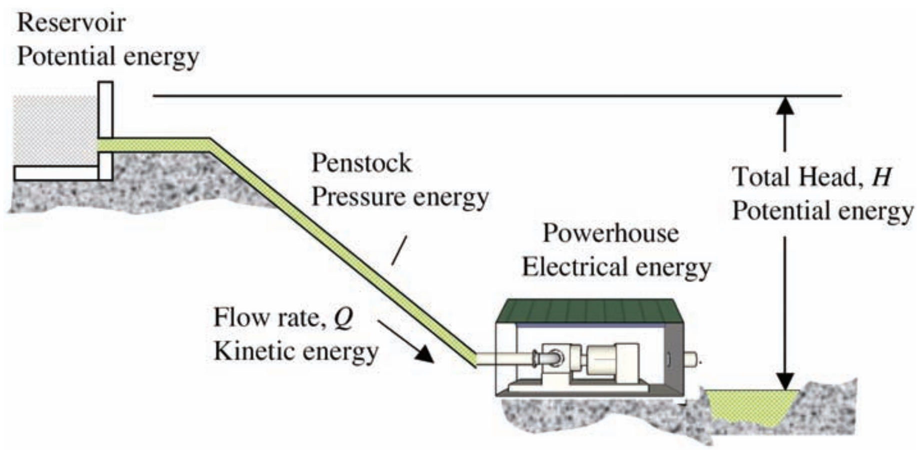
In a month with 30 days, the electric energy delivered would be

$$\text{Energy} = 0.185\ \text{kW} \times 24\ \text{hr/day} \times 30\ \text{day/mo} = 133\ \text{kWh/mo}$$

That's only about one-fourth as much electricity as a small house in the city might use, but this is just a cabin in the woods and you would probably have more than enough to keep your beer cold, play some music, and read well into the night.

Notice the hidden complication in this analysis. Suppose the refrigerator needs 1 kW to start its electric motor, but the generator is supplying only 0.185 kW? There isn't enough oomph to start the motor, and it will likely burn out trying. Your system may have sufficient energy to meet your average demand, but not enough power to meet the peak demand. We could solve this problem, however, by adding some batteries to the system, which could easily supply the peak demand.

figure 4.4



A hydropower system converts energy from potential, to pressure, to kinetic, and then electrical energy.

The hydropower P potentially available from a site is proportional to the difference in elevation between the source and the turbine, called the *head* (H), times the rate at which water flows from one to the other, Q. In Chapter 10, we will work with the following handy relationship between these three quantities:

Eq. 4.8

$$P(\mathrm{W}) = 9810\,\eta\, Q(\mathrm{m}^3/\mathrm{s})\; H(\mathrm{m}) = \frac{\eta\, Q(\mathrm{gpm})\; H(\mathrm{ft})}{5.3}$$

where η = overall efficiency of conversion from potential energy to electrical energy

That efficiency includes the friction losses in the piping, the turbine efficiency, the efficiency of gears or belts to match the rotational speed of the turbine to the needs of the generator, and the efficiency of the generator in converting rotational energy of its shaft to electrical power delivered to the load. Thus, the overall efficiency is

Eq. 4.9

$$\eta = \eta_{\mathrm{penstock}} \cdot \eta_{\mathrm{turbine}} \cdot \eta_{\mathrm{gears}} \cdot \eta_{\mathrm{generator}}$$

Solution Box 4.2 illustrates the use of these equations.

4.3.4 Rotational Energy

Although rotational energy is actually a form of kinetic energy, it is worth considering separately because it is so different from the usual concept of a mass moving along at some linear speed. Old internal combustion engines used to have big, heavy, rotating flywheels to keep the crankshaft turning while they waited for the next burst of exploding fuel. James Watt even had one on his original steam engines. We don't see a lot of these slow-moving flywheels these days, but there is, however, increasing interest in storing energy in high-speed, lightweight flywheels. These new flywheels might replace batteries in such applications as uninterruptible power supplies, backup storage for wind turbines or photovoltaics, or even electric vehicles.

Analogous to Equation 4.7, rotational energy can be written as

Eq. 4.10
$$\text{K.E. (rotational)} = \frac{1}{2} I\omega^2$$

where I = the object's moment of inertia
ω = the object's rotational speed

For a wheel that is rotating, the relationship between its moment of inertia and its mass (m), radius (r), and and shape (k) results in the following, more useful K.E. expression:

Eq. 4.11
$$\text{K.E. (rotational)} = \frac{1}{2} kmr^2\omega^2$$

A bicycle wheel, with essentially all of its mass at the outer perimeter, has $k = 1$, whereas a solid disk of uniform thickness would have $k = 0.5$.

Notice that old flywheels had a lot of mass, but very low rotational speeds. With new composite materials, it is possible to store a lot of energy in a relatively lightweight structure by taking advantage of the fact that energy stored increases as the square of the rotational speed. Thus, for example, a carbon composite flywheel spinning at 20,000 rpm can store 4 million times as much energy in the same mass as an old steel flywheel rotating at only 10 rpm.

Flywheels can be used as the heart of an uninterruptible power supply system. Utility power is used to spin up the flywheel, and utility power normally supplies the load (e.g., your building) directly. In the case of a power outage, energy is extracted from the flywheel, through a generator, to supply power to the load during the outage.

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 were the temperature of a charcoal briquette after a barbeque, we would describe it as being "cool." The context leads us 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). Using 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. 4.12 $$T(°\mathrm{F}) = 1.8\,T(°\mathrm{C}) + 32 \qquad \text{or} \qquad T(°\mathrm{C}) = \frac{5}{9}[T(°\mathrm{F}) - 32]$$

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 that 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 is the same on both the Rankine and Fahrenheit scales. 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. 4.13 $$T(\mathrm{K}) = T(°\mathrm{C}) + 273 \qquad \text{and} \qquad T(\mathrm{R}) = T(°\mathrm{F}) + 460$$

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 referred to as *thermal energy*. Confusing? You bet. The technicality has to do with the "transfer" part of the definition; that is, heat only exists due to 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 its internal energy (U).

Eq. 4.14

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

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), that the net result will be equal to the changes in the internal energy (ΔU), kinetic energy (ΔK.E.) and potential energy (ΔP.E.).

Eq. 4.15

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

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.

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. 4.16

$$\textbf{Change in energy stored} = \Delta E = m\,c\,\Delta T$$

where c = specific heat of the substance

table 4.3 **Specific Heat and Volumetric Heat Capacity of Selected Substances**

	Specific Heat		Density	Heat Capacity
Substance	**(kJ/kg °C)**	**(Btu/lb °F)**	**(lb/ft^3)**	**(Btu/ft^3 °F)**
Water	4.18	1.00	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	0.20	82	16
Gasoline	2.22	0.53	42	22
Steel*	0.46	0.11	487	54

* Representative values

The specific heat is the energy required to raise a unit of mass by one degree. For example, the specific heat of water is

Eq. 4.17 **Specific heat of water c = 1 Btu/lb°F = 4.18 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.

4.5 Chemical Energy

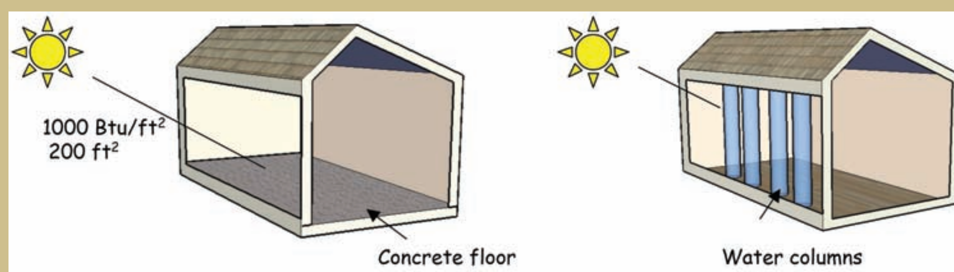
As you recall from basic science, matter is made up of molecules; molecules are made up of atoms; and atoms are made up of protons, neutrons, and electrons. Although physicists are trying to understand even smaller particles, for our purposes we can stop there and just

SOLUTION

SOLUTION BOX 4.3

Storing Heat in Concrete or Water

On a clear, winter day 200 ft^2 of south-facing window allows 1000 Btu/ft^2 of solar energy to hit and be absorbed by some thermal mass in a passive solar house. You want to store that energy during the day and release it at night to help keep your house warm. You could use tubes of water behind the window to store the energy or you could let it be absorbed in a concrete floor slab. What volume of each would be required if the sun heats those masses by 20°F? (Ignore any daytime heat losses from the mass.)



Solution:

The total amount of energy to be stored is 200 ft^2 × 1000 Btu/ft^2 = 200,000 Btu. Using values from Table 4.3, we find

$$\text{Volume of water} = \frac{200{,}000 \text{ Btu}}{62.4 \dfrac{\text{Btu}}{\text{ft}^3 \cdot {}^\circ\text{F}} \times 20^\circ\text{F}} = 160 \text{ ft}^3$$

$$\text{Volume of concrete} = \frac{200{,}000 \text{ Btu}}{30 \dfrac{\text{Btu}}{\text{ft}^3 \cdot {}^\circ\text{F}} \times 20^\circ\text{F}} = 333 \text{ ft}^3$$

So, less than half as much water would be required.

If that concrete were 6 inches thick, it would need to cover 667 ft^2 of floor area. If those water columns were each 8 feet high and 1 foot in diameter, it would take 25 of them.

consider how the energy associated with chemical reactions among atoms and molecules can be captured and utilized.

4.5.1 Atoms and Molecules

You also remember that protons and neutrons form the nucleus of an atom, with the number of protons, called the *atomic number*, identifying which chemical element this is. Protons and neutrons have nearly the same mass (neutrons are very slightly heavier), but the proton carries a positive charge whereas the neutron is electrically neutral. The sum of the number of protons and neutrons is called the *mass number*. Lumped together, the protons and neutrons in a nucleus are referred to as *nucleons*.

Surrounding the nucleus is a swarm (well, just one for hydrogen) of very light, negatively charged electrons, equal in number to the number of positively charged protons, which normally results in an electrically neutral atom. Most of the volume of an atom is empty space. For example, if the orbit of the outer ring of electrons of a typical atom were enlarged to the size of the entire Earth, its nucleus would be only a few hundred feet across.

All atoms with the same chemical name (such as helium) have the same number of protons, but not all such atoms have the same number of neutrons. Elements with the same atomic number but differing mass numbers are called *isotopes*. For example, helium always has 2 protons (by definition), but it may have anywhere from 1 to 8 neutrons. Almost always it has only 1 neutron (Helium-3), roughly one in a million helium atoms has two neutrons (Helium-4), while isotopes with more than 2 neutrons are unstable and occur only for short periods of time during nuclear reactions.

The most common way to describe a given isotope is by giving its chemical symbol with the mass number written at the upper left and the atomic number at the lower left. For example, the two most important isotopes of uranium (which has 92 protons) are

$$\underset{\textbf{Uranium-235}}{{}^{235}_{92}\mathrm{U}} \qquad \textbf{and} \qquad \underset{\textbf{Uranium-238}}{{}^{238}_{92}\mathrm{U}}$$

When referring to a particular element, it is common to drop the atomic number subscript because it adds little information to the chemical symbol. Thus “U-238” and “Helium-3” are common ways to describe those isotopes.

When atoms are close enough together that their outer electrons can interact with each other, forces can be set up between those atoms that are strong enough to hold them together. Nearby atoms share outer electrons forming *covalent bonds*, with each of a pair of atoms providing an electron that the other shares. *Molecules*, then, are made up of atoms held together by these covalent bonds.

4.5.2 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. These molecules are said to be in *solid, crystalline state*.

When those molecules are heated, 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 required 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 (see Solution Box 4.4).

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 is 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 required 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).

4.5.3 Stoichiometry: Mass Balance in Chemical Reactions

The principle of conservation of mass can be applied to a chemical reaction to tell us how much of each compound is involved to produce the results shown. The balancing of equations so that the same number of each kind of atom appears on each side of the equation, and the subsequent calculations to determine amounts of each compound involved, is known as *stoichiometry*.

Consider the following simple reaction in which methane (CH_4) is oxidized (burned) to produce carbon dioxide and water. Because a significant fraction of our energy comes from natural gas, and because natural gas is mostly methane, this reaction is extremely important.

Eq. 4.18
$$CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O$$

Let's use Equation 4.18 to help us review a little bit of chemistry. First, notice the reaction as written is balanced; that is, there are as many atoms of carbon, hydrogen, and oxygen on the left side of the reaction as there are on the right.

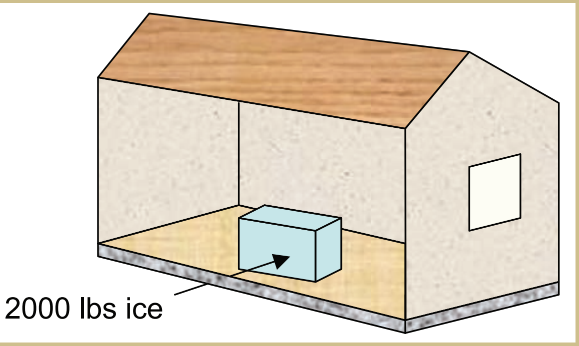
Now, we can interpret the equation to say that 1 molecule of methane reacts with 2 molecules of oxygen to produce 1 molecule of carbon dioxide plus 2 molecules of water. It is of more use, however, to be able to describe this reaction in terms of the mass of each substance (for example, How many grams of CO_2 are put into the atmosphere when so many

SOLUTION BOX 4.4

Cooling Your Room with a Ton of Ice

SOLUTION

Suppose you want to cool your room by dragging in a ton of ice (2000 lb) and letting it melt. How much heat would the ice need to take from the room to cause it all to melt? If it took 24 hours for that to happen, what average rate of cooling would you have achieved?



Solution:

We have learned that the latent heat of fusion for water is 144 Btu/lb, so the amount of heat needed to melt 1 ton of ice would be

$$\mathbf{2000\ lb \times 144\ Btu/lb = 288{,}800\ Btu}$$

If it takes 24 hours to melt that ton of ice, the average rate of heat removal from the room, and hence cooling of the room, would be

$$\mathbf{\frac{288{,}800\ Btu}{24\ hr} = 12{,}000\ Btu/hr}$$

As it turns out, the standard way to rate the cooling capacity of an air conditioner in the United States is by the tons of cooling it can provide, where 1 ton = 12,000 Btu/hr. Thus, for example, a 3-ton home air conditioner provides 36,000 Btu/hr of cooling, equivalent to melting 3 tons of ice per day.

grams of CH_4 are burned?, and so forth). To do so, we need to define the atomic weight and molecular weight of atoms and molecules, and we need to lump large numbers of molecules into chunks called *moles*.

The *atomic weight* of an element is expressed in *atomic mass units* (amu), where one amu is defined to be exactly one-twelfth the mass of carbon-12. (C-12 has 6 protons and 6 neutrons.) Why bother measuring mass in amu's rather than just expressing it with the mass number (total protons plus neutrons)? To answer that, recall that chemical elements often have various isotopes in nature—carbon, for example, comes with 6 neutrons in most atoms (C-12), but some have 8 neutrons (C-14). That suggests the naturally occurring mix of carbon should have an atomic weight a bit higher than 12. Because the mass number and the atomic weights of an element are so similar, we will henceforth ignore that distinction and follow the somewhat standard engineering practice of rounding off the atomic weights (e.g., C will be 12 amu rather than 12.011).

The molecular weight of a molecule is simply the sum of the atomic weights of the constituent atoms. Thus, the (slightly rounded) molecular weight of CH_4 is $12 + 4 \times 1 = 16$ amu. If we divide the mass of a substance by its molecular weight the result is the mass expressed in moles (mol). For example, 32 grams of CH_4, divided by 16 g/mol, is 2 moles of methane.

The special advantage of expressing chemical reactions using moles is that one mole of any substance contains exactly the same number of molecules (1 g-mole = 6.02×10^{23} molecules), which gives us another way to interpret a chemical reaction such as that given in Equation 4.18:

$$\begin{array}{ccccccc} CH_4 & + & 2\,O_2 & \rightarrow & CO_2 & + & 2\,H_2O \\ 1\text{ mol } CH_4 & + & 2\text{ mol } O_2 & \rightarrow & 1\text{ mol } CO_2 & + & 2\text{ mol } H_2O \end{array}$$

We can now describe the reaction by saying that one mole of CH_4 reacts with two moles of O_2 to produce one mole of CO_2 and 2 moles of H_2O. We can translate that into grams by first converting each constituent into moles using their (rounded) atomic weights:

$$\begin{aligned} CH_4 &= 12 + 4 \times 1 = 16 \text{ g/mol} \\ O_2 &= 2 \times 16 = 32 \text{ g/mol} \\ CO_2 &= 12 + 2 \times 16 = 44 \text{ g/mol} \\ H_2O &= 2 \times 1 + 16 = 18 \text{ g/mol} \end{aligned}$$

We now have a third way of expressing the oxidation of methane:

$$\begin{array}{ccccccc} CH_4 & + & 2\,O_2 & \rightarrow & CO_2 & + & 2\,H_2O \\ 16\text{ g } CH_4 & + & 64\text{ g } O_2 & \rightarrow & 44\text{ g } CO_2 & + & 36\text{ g } H_2O \end{array}$$

Notice that mass is conserved in this last expression; that is, there are 80 grams on the left and 80 grams on the right.

4.5.4 Enthalpy: The Energy Side of Chemical Reactions

Most of the energy we use to power our industrial societies is obtained by burning fossil fuels; mainly, coal, oil, and natural gas. Intuitively, we know that burning a fuel converts energy stored in chemical bonds into heat that we can use to do work. Chemical reactions in which heat is liberated, such as those that occur when fuel is burned, are called *exothermic* reactions. Reactions that want to go in the other direction; that is, in which heat is required to make them happen, are called *endothermic* reactions.

Just as we use stoichiometry to do mass balances on chemical reactions we can use something called *enthalpy* to help us do energy balances. As is often the case with thermodynamic properties of substances, the precise definition of enthalpy is rather subtle and does not lend itself to simple interpretation. One way to think about it, however, is that it is a measure of the energy that it takes to form a substance out of its constituent elements, in which case it is called the *enthalpy of formation*.

Table 4.4 lists some enthalpies of formation for selected substances. Notice that gaseous oxygen (O_2) and hydrogen (H_2) are listed as having zero enthalpy. Enthalpy needs a reference point, just as other forms of energy do (think potential energy), and the zero point on the enthalpy scale is applied to the stable form of a chemical element under standard temperature and pressure (STP) conditions (25°C and 1 atmosphere of pressure). Because, for example, the stable form of oxygen under STP conditions is molecular O_2, it is given an enthalpy value of zero, but atomic oxygen O, which is not stable, does not have zero enthalpy.

Also notice that enthalpies depend on the state of the substance; that is, the enthalpy of liquid water is different from that of gaseous water vapor. In the case of H_2O, that difference is very important, as we shall see later.

In chemical reactions, the difference between the enthalpies of the products (right side) and the reactants (left side) tells us how much energy is released or absorbed in the reaction. When there is less enthalpy in the final products than in the reactants, heat is liberated; that is, the reaction is exothermic.

By combining stoichiometric and enthalpy analyses of a chemical reaction we can, with modest effort, determine such important characteristics of a fuel as the energy and carbon released when the fuel is burned. This is demonstrated in Solution Box 4.5.

table 4.4 Enthalpy of Formation for Selected Substances (STP)

Substance	Formula	State	Enthalpy, *H* (kJ/mol)
Hydrogen	H_2	Gas	0
Oxygen	O	Gas	247.5
Oxygen	O_2	Gas	0
Water	H_2O	Liquid	–285.8
Water vapor	H_2O	Gas	–241.8
Methane	CH_4	Gas	–74.9
Carbon dioxide	CO_2	Gas	–393.5
Methanol	CH_3OH	Liquid	–238.7
Ethanol	C_2H_5OH	Liquid	–277.7
Propane	C_3H_8	Gas	–103.9
Octane	C_8H_{18}	Liquid	–250.0
Glucose	$C_6H_{12}O_6$	Solid	–1260

SOLUTION BOX 4.5

Heat and Carbon Emissions from Combustion of Methane

To demonstrate the use of enthalpy, let us return to the combustion of methane given in Equation 4.18 but now let us write the associated enthalpies under each substance as given in Table 4.4. Notice we have to decide whether the water released will be in the vapor form or liquid. In this case, we assume it is vapor (later we'll see that this means our result will be what is called the lower heating value, LHV, of methane).

$$\begin{array}{ccccccc} \mathbf{CH_4} & + & \mathbf{2\,O_2} & \rightarrow & \mathbf{CO_2} & + & \mathbf{2\,H_2O(g)} \\ \mathbf{(-74.9)} & & \mathbf{2 \times (0)} & & \mathbf{(-393.5)} & & \mathbf{2 \times (-241.8)} \end{array}$$

The change in enthalpy associated with this reaction is then

$$\Delta H = (\Sigma\ H \text{ of products}) - (\Sigma\ H \text{ of reactants})$$
$$\Delta H = [-393.5 + 2 \times (-241.8)] - [(-74.9) + 2 \times (0)] = -802.2 \text{ kJ/mol of } CH_4$$

Because the enthalpy change is negative, that means the reaction is exothermic. It releases 802.2 kJ of heat for every mole (16 g) of methane burned; that is, 50.14 kJ/g.

To determine carbon emissions, we note that one mole of CH_4 (16 g) produces one mole of CO_2 (44 g). Combining that with the 50.14 kJ of energy released per gram of CH_4 gives

$$\frac{44 \text{ g } CO_2}{12 \text{ g } CH_4} \times \frac{\text{g } CH_4}{50.14 \text{ kJ}} = 0.05485 \text{ g } CO_2/\text{kJ} = 54.85 \text{ kg } CO_2/\text{MJ}$$

Quite often, emissions are expressed in kg of C (not CO_2) per MJ of energy. That's easy to fix. Because CO_2 has 12 g C per 44 g CO_2, we can adjust the emission rate to be

$$\frac{54.85 \text{ kg } CO_2}{\text{MJ}} \times \frac{12 \text{ kg C}}{44 \text{ kg } CO_2} = 15.0 \text{ kg C/MJ}$$

4.5.5 Heat of Combustion: HHV and LHV

When a fuel is burned, the magnitude of the enthalpy difference $\Delta(H)$ between the reactants and the products is called the *heat of combustion* (or, sometimes, *heat content*) of the fuel. Table 4.5 presents some values for various commonly used fuels, expressed in typical U.S. units.

Notice that Table 4.5 qualifies the heat content of those fuels by saying it is the HHV value. Also given is an HHV to LHV ratio. What does that mean?

table 4.5 Heat of Combustion for Selected Fuels (Average HHV Values)

Fuel	Heat of Combustion (HHV)	HHV/LHV
Coal (bituminous)	14,000 Btu/lb	1.05
Fuel ethanol	84,262 Btu/gallon	1.11
Fuel oil (#2)	140,000 Btu/gallon	1.06
Gasoline	125,000 Btu/gallon	1.07
Hydrogen	61,400 Btu/lb	1.18
Kerosene	135,000 Btu/gallon	1.06
Methanol	64,600 Btu/gallon	1.14
Natural gas	1025 Btu/cu. ft.	1.11
Pellets (for pellet stove; premium)	8250 Btu/lb	1.11
Propane	91,330 Btu/gallon	1.09
Wood (20% moisture)	7000 Btu/lb	1.14

* For comparison, the ratios of HHV to LHV are given as well.

Recall from section 4.5.2 that it takes quite a lot of energy, called the latent heat of vaporization, to cause water to change state from a liquid to a gas. When a fossil fuel is burned, some of its energy ends up as latent heat in the water vapor produced. Whether the latent heat is, or is not, included leads to two different values of the heat of combustion. The higher heating value HHV, also known as the *gross* heat of combustion, includes the latent heat of the water vapor. That is, HHV includes all of the heat that could possibly be captured if you could condense the water vapor and utilize it for your process. The lower heating value (LHV), also known as the *net* heat of combustion, does not include latent heat. For example, the 802.2 kJ/mol calculated for the combustion of methane in Solution Box 4.5 didn't include condensing the water vapor, so that is the LHV for methane. If we were to repeat that example using the enthalpy of liquid water, an HHV of 890.2 kJ/g for methane would have been found.

Solution Box 4.6 illustrates the enthalpy calculations that lead to both the HHV and LHV of hydrogen. Notice the difference between the LHV and HHV values is about 18%, which is pretty sizeable. For comparison, the difference between LHV and HHV for natural gas is about only 11%, whereas for coal it is less than 6%. In general, the simpler the fuel molecule—that is, how little carbon it contains—the more important it becomes to be careful about distinguishing between LHV and HHV.

When calculating efficiencies of various processes, it is important to specify whether the answers obtained are based on LHV or HHV—especially when there are significant differences between the two systems, as illustrated in Solution Box 4.7.

Unfortunately, the distinction between LHV and HHV often leads to considerable confusion. If you know what LHV and HHV mean, and it is not mentioned in someone's description of an energy transformation, you could be somewhat uncertain about how to interpret that person's results. On the other hand, if it is mentioned and you don't know anything about

SOLUTION

SOLUTION BOX 4.6

HHV and LHV for Hydrogen

The oxidation of hydrogen, either by burning, or perhaps through a chemical reaction in a fuel cell, can be described as follows:

$$\textbf{LHV} \qquad H_2 + \frac{1}{2}O_2 \rightarrow H_2O \text{ (gas)} \qquad \Delta H = -241.8 \text{ kJ/mol}$$

$$\textbf{HHV} \qquad H_2 + \frac{1}{2}O_2 \rightarrow H_2O \text{ (liquid)} \qquad \Delta H = -285.8 \text{ kJ/mol}$$

Notice how easy it was to write those reactions. Because the enthalpy of formation of H_2 and of O_2 is zero, the overall energy released is just the numerical value of the enthalpy of formation of water vapor (for LHV) or liquid water (for HHV).

Because there are 2 g per mole of hydrogen, its LHV is 241.8/2 = 120.9 kJ/g and its HHV is 285.8/2 = 142.4 kJ/g.

LHV and HHV, you again could be left scratching your head. To compound the problem, all energy data given the Energy Information Administration's *Annual Energy Reviews* (the principal source in the United States) are HHV values, whereas most of the rest of the world uses LHV as its standard. Moreover, even in the United States, for small power plants—fuel cells, microturbines, and the like—LHV is often used. So, which makes more sense?

The argument many make for LHV values is that the latent heat in the water vapor produced during combustion is almost always lost out the stack, along with the other combustion gases, so, it shouldn't be counted. On the other hand, there are, for example, very efficient home furnaces these days that capture latent heat by purposely cooling exhaust gases enough to cause water vapor to condense. On an HHV basis, these condensing furnaces can have efficiencies of well over 90%. If we were to use the LHV basis for the fuel, such furnaces can have an efficiency of over 100%, which sounds a little strange even though it is correct.

The bottom line is that for casual computations, in most circumstances the difference between LHV and HHV is only a few percent and can usually be ignored. When more precision is required the most important thing is to be consistent and base everything on either LHV or HHV, and be careful to specify which system you are using.

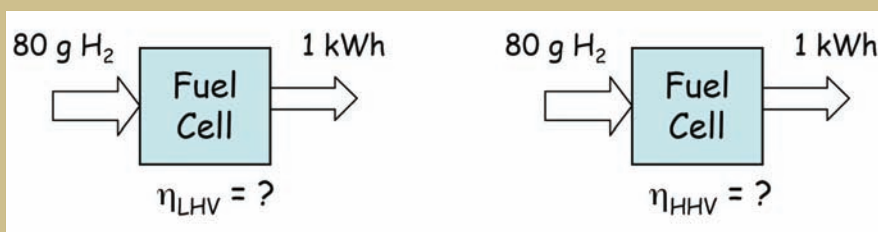
4.6 Solar Energy

The source of energy that keeps our planet at just the right temperature, powers our hydrologic cycle, creates our wind and weather, and provides our food and fiber is a relatively insignificant yellow dwarf star some 93 million miles away. Powering our sun are thermonu-

SOLUTION BOX 4.7

LHV and HHV Efficiency of a Fuel Cell

Suppose a fuel cell converts 80 g of hydrogen into 1 kWh (3600 kJ) of electricity. Using the lower heating value (LHV) for H_2, what is the energy efficiency of the fuel cell? What is its efficiency on an HHV basis?



Solution: Using the results found in Solution Box 4.6, 80 g of hydrogen can provide 120.9 kJ/g × 80 g = 9672 kJ on an LHV basis (no condensation). The energy delivered by the fuel cell is 1 kWh, which is 3600 kJ (Table 4.1). So, on an LHV basis the fuel-cell efficiency would be

LHV efficiency = 3600 kJ/9672 kJ = 0.372 = 37.2%

On an HHV basis, 80 g of H_2 can provide 142.4 kJ/g × 80 g = 11,392 kJ. So the HHV efficiency is

HHV efficiency = 3600 kJ/11,392 kJ = 0.316 = 31.6%

Thus, the same fuel cell, producing the same output from the same amount of hydrogen, could be described as having an efficiency of either 37.2% or 31.6% depending on whether efficiency is determined using an LHV or HHV basis. Most often, fuel-cell efficiencies are, perhaps not surprisingly, stated using the LHV basis.

clear reactions in which hydrogen atoms fuse together to form helium. In the process, about 4 billion kilograms of mass per second are converted into energy following Einstein's famous relationship $E = mc^2$. This fusion has been continuing reliably for the past 4 or 5 billion years and is expected to continue for another 4 or 5 billion.

At the center of the sun, temperatures are estimated to be on the order of 15 million K, but those incredibly high temperatures are degraded as energy works its way some 400,000 miles from the core to the surface. The surface of the sun radiates about 3.8×10^{20} MW of electromagnetic energy outward toward space with wavelengths that correspond closely to what would be expected from an object around 5800 K.

Just outside of our atmosphere, the Earth receives an average of about 1.35 kW/m^2 of radiant energy. Some of that energy is reflected back into space and some is absorbed in the atmosphere, and of course half of the time at every spot on Earth there is no sunlight at all. The result is that an average of about 160 watts per square meter of solar energy actually reaches the Earth's surface. To put that into perspective, the sunlight falling onto roads in just the United States alone is equivalent to the entire global rate of consumption of fossil fuels. And, some claim that no country on Earth uses as much energy for all purposes as is contained in the sunlight striking its buildings.

Subsequent chapters will explore the conversion of sunlight into useful forms of energy, but for now we will just examine the characteristics of electromagnetic energy itself.

4.6.1 Electromagnetic Radiation

Electromagnetic radiation can be described in terms of discrete zero-mass particles of energy, called photons, or in terms of electromagnetic waves of various wavelengths and frequencies. Both are correct, and the descriptor to use is mostly a matter of which one is more convenient for the particular phenomenon being described. The origin of electromagnetic radiation is the motion of electrically charged particles. It occurs when atoms collide with each other or when absorption of incoming photons causes atoms to temporarily transition to higher-energy, unstable states. As excited atoms relax, photons are released. Radiation can also be the result of nuclear decay reactions or other nuclear and subnuclear processes. Electromagnetic radiation can travel through empty space, or through air or other substances.

When described as a wave phenomenon, such as is shown in Figure 4.5, the wavelength and frequency of vibration of electromagnetic radiation are related as follows:

Eq. 4.19
$$\lambda = \frac{c}{\nu}$$

where λ = wavelength (m)
c = speed of light (3×10^8 m/s)
ν = frequency (hertz, i.e., cycles per second)

When radiant energy is described in terms of photons, the relationship between frequency and energy is given by

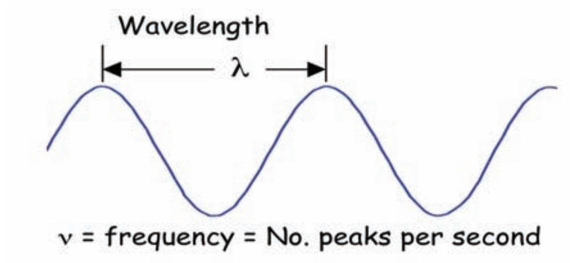
Eq. 4.20
$$E = h\nu$$

where E = energy of a photon (J)
h = Planck's constant (6.6×10^{-34} J-s)

Equation 4.20 indicates that photons with higher frequency (shorter wavelengths) have higher energy.

figure 4.5

Electromagnetic radiation as a wave phenomenon is described in terms of frequency and wavelength.

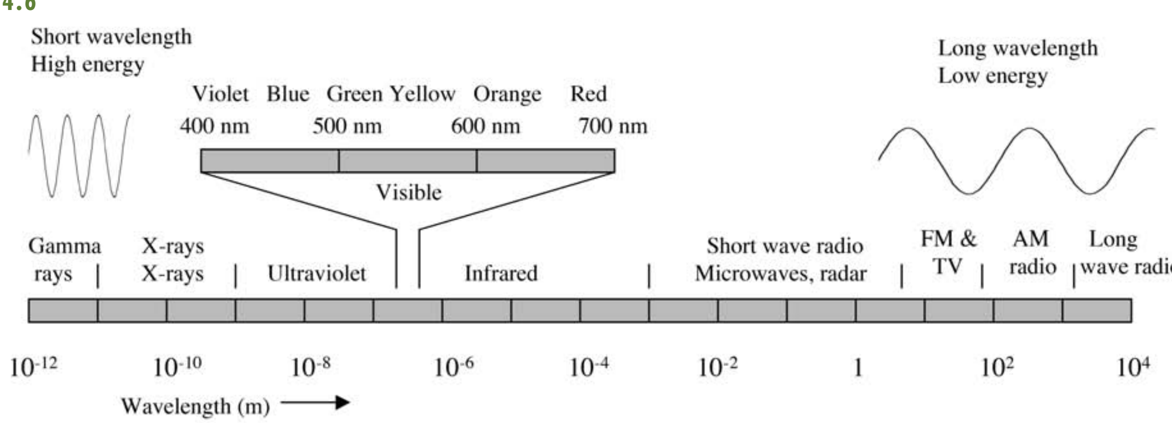


Radio waves, microwaves, visible light, and x-rays are all examples of electromagnetic radiation. Each is characterized by wavelengths that range over some portion of the electromagnetic spectrum. As shown in Figure 4.6, the range extends from long-wave radio waves, with wavelengths thousands of meters long, to gamma rays, with wavelengths down in the vicinity of 10^{-12} m. The solar portion covers only a small fraction of these wavelengths, 200 to 2500 nanometers (1 nm = 10^{-9} m), with the visible portion—the part that enables us to see things—extending from about 400 to 700 nm.

4.6.2 The Solar Spectrum

Every object emits thermal radiation, the characteristics of which depend on the object's temperature. The usual way to describe how much radiant energy the object emits, as well as the characteristic wavelengths of that electromagnetic energy, is to compare it to a theoretical abstraction called a *blackbody*. A blackbody is defined to be a perfect emitter and a perfect absorber. As a perfect emitter, it radiates more energy per unit of surface area than any real object at the same temperature. As a perfect absorber, it absorbs all radiant energy that strikes its surface; that is, none is reflected and none is transmitted through it.

figure 4.6 **The Electromagnetic Spectrum**



The total rate at which energy is emitted from a blackbody with surface area A and absolute temperature T is given by the *Stefan-Boltzmann law of radiation*:

Eq. 4.21

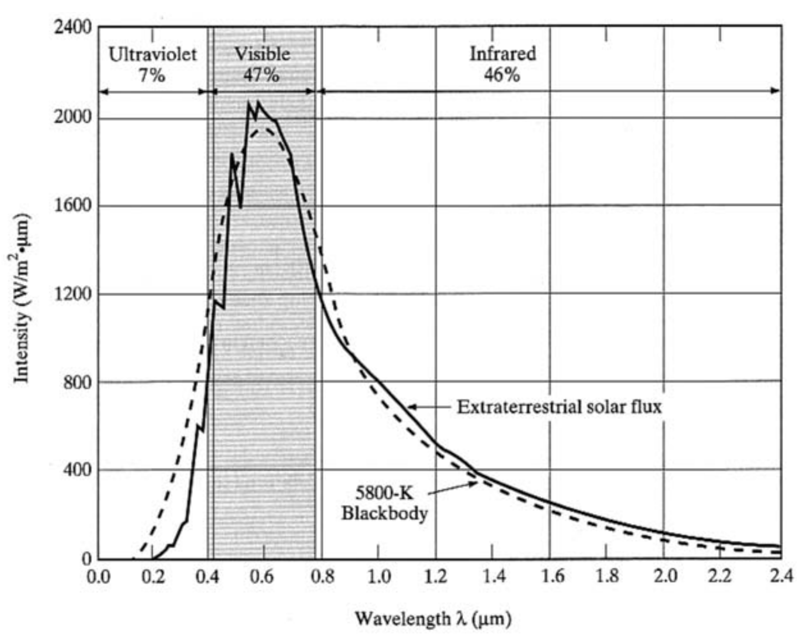
$$E = \sigma A T^4$$

where E = total blackbody emission rate (W)
σ = the Stefan-Boltzmann constant = 5.67×10^{-8} W/m^2 – K^4
T = absolute temperature (kelvins K, where K = 273 + °C)
A = surface area of the object (m^2)

Actual objects do not emit as much radiation as our hypothetical blackbody, but most are quite close to this theoretical limit. The ratio of the amount of radiation an actual object emits, to that of a blackbody, is called the object's *emissivity*. The emissivity of desert sand, dry ground, and most woodlands is estimated to be 0.90, whereas water, wet sand, and ice have an emissivity of roughly 0.95. A human body, no matter what pigmentation, has an emissivity of around 0.96.

Although Equation 4.21 gives us the total emissive power of a blackbody, it doesn't describe the range of wavelengths associated with that radiation. A more complex equation, known as *Planck's law*, provides that spectral distribution of energy, an example of which is shown in Figure 4.7. In that figure, the spectrum of a 5800 K blackbody is compared with the actual spectrum of solar energy arriving just outside of Earth's atmosphere. The closeness

figure 4.7



The extraterrestrial solar spectrum (solid line) compared with the spectrum of a 5800 K blackbody (dashed). Also shown is the fraction of solar energy that falls with the ultraviolet, visible, and infrared portions of the solar spectrum.

of the match demonstrates quite clearly that the sun can quite reasonably be mathematically modeled as if it were a 5800 K blackbody.

The area under a spectral distribution curve between any two wavelengths is the total radiant power within that region. In Figure 4.7, the solar spectrum is divided into three regions: ultraviolet, visible, and infrared. Roughly half of the energy in the extraterrestrial solar spectrum (47%) is contained within the visible wavelengths (0.38 to 0.78 μm). The ultraviolet (UV) is only 7%, but because those are such short wavelength photons they pack a more energetic punch and can be especially damaging to living things. Fortunately, stratospheric ozone filters most of those dangerous wavelengths out of the spectrum before they reach the Earth's surface (unless you happen to live under the seasonal ozone hole). The infrared portion makes up 46% of the total. These photons help keep us warm, but they don't contribute to our ability to see anything. Later, when we describe energy-efficient building technologies, we will see that window coatings for office buildings can be designed to transmit just the visible wavelengths for natural daylighting, while blocking out the UV to help keep fabrics from fading, and reflecting the infrared to help reduce the air-conditioning load.

4.6.3 The Greenhouse Effect

The wavelength at which the spectrum peaks is sometimes a handy way to characterize blackbody radiation. Objects with higher temperatures have their peak at a shorter wavelength as described by *Wien's displacement rule*:

Eq. 4.22

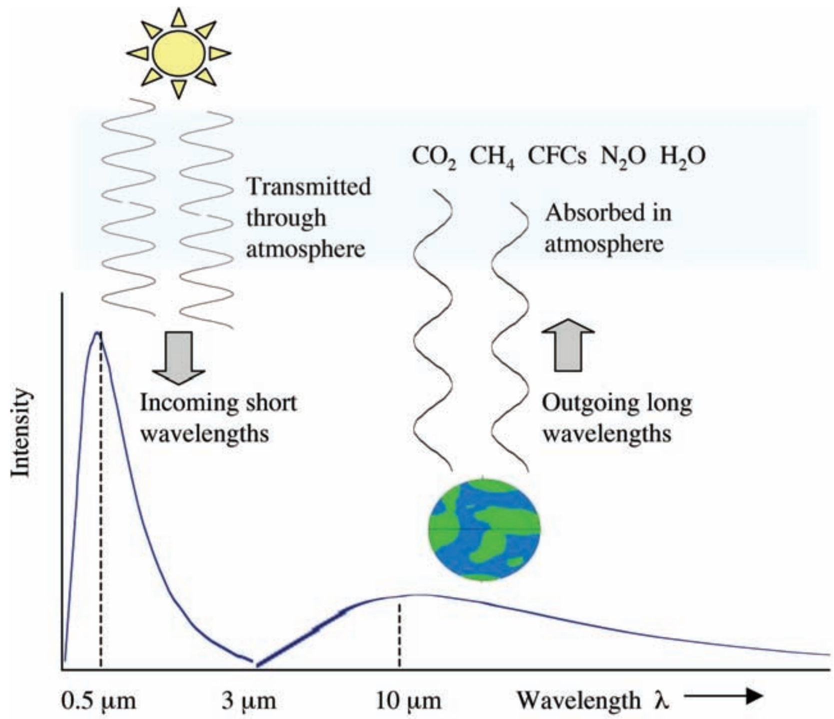
$$\lambda_{max}\,(\mu\text{m}) = \frac{2898}{T(\text{K})}$$

This equation predicts that the sun, at 5800 K, has a peak at about 0.5 μm, which agrees with Figure 4.7. The Earth, with its surface at about 15°C (298 K), should show its spectral peak at about 10 μm, which it does (Figure 4.8).

That large difference between incoming solar wavelengths (short) and outgoing wavelengths radiated to space from the Earth's surface (long) is critical to understanding the greenhouse effect. As it turns out, our atmosphere is relatively transparent to the short wavelengths coming from the hot sun, but it is relatively opaque to longer infrared wavelengths trying to work their way from the Earth's surface, through the atmosphere, and back to outer space. By preferentially absorbing outgoing radiation, greenhouse gases, including CO_2, CH_4, N_2O and water vapor, act like an insulating blanket around the Earth causing it to be about 19°C warmer than it would be without those gases. As Solution Box 4.8 demonstrates, without the naturally occurring greenhouse effect, the Earth would likely be a virtually uninhabitable frozen planet. Our current worry about global warming is of course based on human interactions that are increasing the greenhouse effect, with uncertain, but potentially devastating, consequences.

The calculation in Solution Box 4.8 indicates that without the greenhouse effect the average temperature of the Earth would be about 34°C below its actual 15°C temperature. That is, the Earth would be well below zero and all water would be frozen solid.

figure 4.8



Incoming solar wavelengths more easily pass through the atmosphere than outgoing long-wave radiation from the Earth's surface.

4.6.4 Solar Energy for Living Things

Although the emphasis in this book is on capturing, converting, and using energy for material things in life (hot showers and cold beer), without the energy services that nature freely provides for us life as we know it today would not exist. Sunlight provides the foundation for the biosphere. Green plants bottle up sunlight in energy-rich bonds using just water and carbon dioxide as raw materials, from which they build sugars that are the energy basis for essentially all life on Earth. And, as a bonus, they pump fresh oxygen into the atmosphere. The same chlorophyll photon-trapping processes that enable life to thrive on Earth today are responsible for the ancient solar energy, bottled up hundreds of millions of years ago, that we now exploit in the form of fossil fuels.

The process of photosynthesis utilizes chlorophyll, the green pigment in the chloroplasts of plants and some algae. Chlorophyll absorbs sunlight from the red portion of the solar spectrum (around 0.7 micron) whereas other pigments absorb shorter-wavelength blue light. Note that green light is not absorbed, and is instead reflected, which is why the leaves of plants appear green to us because those green wavelengths are reflected to our eyes.

A simple representation of photosynthesis is the following in which photons from the sun provide the energy needed to convert simple water and CO_2 molecules into sugars, in this case glucose, $C_6H_{12}O_6$, with the leftover oxygen being released into the atmosphere.

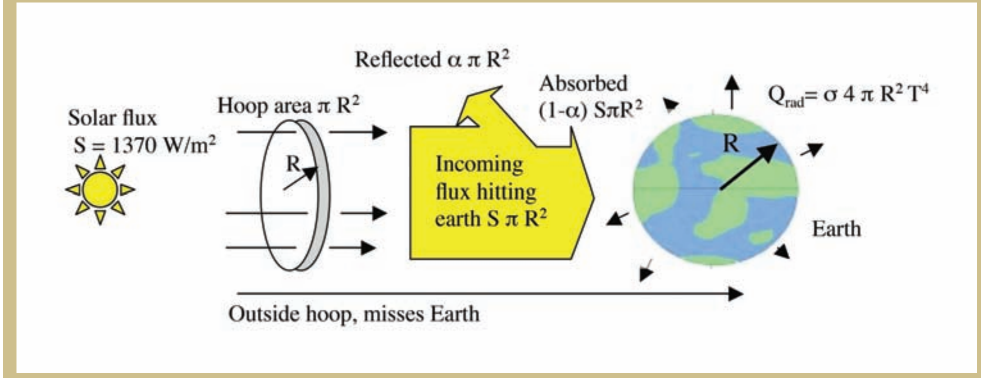
SOLUTION BOX 4.8

Estimating the Earth's Temperature without the Greenhouse Effect

Let's do a calculation for the expected temperature of the Earth if it were a blackbody without an atmosphere, and hence, without any greenhouse effect. We will find the equilibrium temperature that would cause the energy radiated to space from the Earth's surface to be equal to the solar energy absorbed by the surface.

Solution:

Imagine a transparent hoop located between the sun and the Earth. The Earth and the hoop have the same radius R and the hoop is set up so that any solar energy (S, W/m^2) that passes



through the hoop hits the Earth and any that misses the hoop also misses the Earth. Some fraction of incoming solar radiation, called the *albedo* (α), is reflected directly back into space.

Incoming Solar Energy:

Solar flux that passes through the hoop and hits the Earth = $S \pi R^2$

Flux hitting the Earth that is reflected = $\alpha S \pi R^2$

Flux hitting the Earth that is absorbed by the Earth = $(1 - \alpha) S \pi R^2$

Outgoing Energy Radiated from Earth:

Assume the Earth is a blackbody with uniform surface temperature = T

Area of Earth's surface = $4 \pi R^2$

Radiation from Earth toward space = $\sigma A T^4 = \sigma 4 \pi R^2 T^4$

(continued on next page)

SOLUTION

Energy Balance:

Incoming solar energy = Outgoing energy radiated from Earth

$$(1 - \alpha)\, S\, \pi\, R^2 = \sigma\, 4\, \pi\, R^2\, T^4$$

The Earth's α is about 0.31 (31%) and the incoming solar flux just outside of the atmosphere, called the solar constant S, is about 1370 W/m^2. Solving the above equation for the equilibrium temperature for Earth gives:

$$T = \left[\frac{S(1-\alpha)}{4\sigma}\right]^{1/4} = \left[\frac{1370\ \text{W/m}^2\ (1-0.31)}{4 \times 5.67 \times 10^{-8}\ \text{W/m}^2\text{K}^4}\right]^{1/4} = 254\ \text{K} = -19°\text{C} = -2°\text{F}$$

Eq. 4.23 $$6\ CO_2 + 6\ H_2O + \text{light} \rightarrow C_6H_{12}O_6 + 6\ O_2 \quad \Delta H = 2820\ \text{kJ/mol glucose}$$

As indicated, the enthalpy change is positive, meaning that 2820 kJ of energy per mole of glucose must be provided to make this reaction take place. At 180 g/mol that works out to be 15.7 kJ/g stored. The annual capture of solar energy by plants is enormous, something like 3000 exajoules (3000×10^{18} J/yr), which is almost ten times the rate of human energy use. In Chapter 14, the use of some of this energy in the form of biomass fuels will be explored.

The overall efficiency at which solar energy is converted to ecosystem biomass varies widely. Some estimates suggest the theoretical maximum to be about 5%, but actual ecosystems are usually far below that value. Whereas a healthy cornfield may capture 1%–2% of the summer's sunlight, a prairie may net only a few tenths of one percent. Globally, the fraction of sunshine that ends up as stored energy in plants is just a few hundredths of a percent.

The reverse of Equation 4.23 summarizes the respiratory process by which living things use stored chemical energy to provide for their own energy needs.

Eq. 4.24 $$C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O \qquad \Delta H = -2820\ \text{kJ/mol}$$

Respiration uses up some of the energy that plants capture during photosynthesis, resulting in less energy available for animals that eat those plants. As we work our way up the food chain, the fraction of consumed energy that is needed to meet the respiratory needs of the next trophic level increases. A common rule-of-thumb estimate is that every time we move one notch up the food chain, from plant to herbivore, from herbivore to first-level carnivore, and so on, only about 10% of one level's energy is transferred to the next. This argument can be restated to suggest that it takes about ten times the land area to feed carnivores than herbivores, which is an often-cited reason why some people choose to be vegetarians.

The magic of both photosynthesis and respiration is provided by a number of enzymes and nucleotides, such as adenosine diphosphate (ADP) and adenosine triphosphate (ATP), that transport and accept electrons, permitting the chemical reactions for biosynthesis of

table 4.6 Examples of Food Calories (kcal)

Food	Calories	Food	Calories
Pizza (3.2 oz)	260	Granola bar	120
Big Mac hamburger	560	Yogurt, nonfat (6 oz)	100
Milkshake (10 oz)	350	Milk, nonfat (8 oz)	85
French fries (large)	400	Cheerios (1¼ cup)	110
Oat bran muffin	330	Carrot, raw	30
Cake donut	270	Lettuce, iceburg (1 head)	70
Twinkies, each	160	Bagel, plain	150
Pecan pie (1/8 of 9-in.)	430	Bread, whole wheat (1 slice)	70
Peanut butter cups, each	140	Apple, medium	70
Ice cream, vanilla soft	380	Halibut, broiled (4 oz)	195

DNA and proteins, assembly of cell structures, transport of solute molecules, neurological information transfer by the nerves and senses, muscle contraction and motion, and all the other physical miracles of life.

4.6.5 Food Calories

Sunshine becomes the calories we eat, so let's take a brief look at food. The nutritional community traditionally uses *Calories* (capital C) to indicate the metabolizable energy content of our food, where 1 Calorie = 1 kilocalorie = 4.185 kJ. Often they don't stick with the convention of using a capital C, so just realize when someone talks about food calories they really mean kilocalories. Table 4.6 provides some examples of the Calories in typical foods we eat.

Table 4.7 provides some estimates of the rate at which an adult will burn Calories while performing various activities. The baseline for a body at rest is called the basal metabolism rate, which corresponds to the energy needed just to provide basic functions such as breathing and blood circulation. Also shown is the equivalent amount of heat given off by a 70 kg (154 lb) individual, expressed in watts. For example, a 154 lb adult watching TV is equivalent to an 86-W heater sitting in the room. Later, when we look at the heating and cooling requirements for a building we will see that we need to account for the heating load provided by people and appliances.

Combining Tables 4.6 and 4.7 allows us to do simple calculations that are rather illuminating. For example, we might wonder how long it would take for a 180 lb man to jog off that Big Mac and milkshake that he had for lunch.

$$\text{Calories} = 560 + 350 = 910 \text{ kcal}$$
$$\text{Jogging} = 3.61 \text{ kcal/hr/lb} \times 180 \text{ lb} = 650 \text{ kcal/hr}$$
$$\text{Time} = 910 \text{ kcal}/650 \text{ (kcal/hr)} = 1.4 \text{ hr}$$

table 4.7 Approximate Calories Used by an Adult, per Unit of Body Weight

Activity	kcal/hr per kg	kcal/hr per lb	Watts per 70 kg (154 lb)
Basal metabolism	1.06	0.48	86
Watching TV	1.06	0.48	86
Driving a car	2.65	1.20	215
Swimming (slow)	4.24	1.93	345
Walking (4 mph)	6.35	2.89	517
Jogging (5 mph)	7.94	3.61	646
Fast dancing	8.82	4.01	718
Bicycling (13 mph)	9.40	4.27	765
Swimming (fast)	12.65	5.75	1030
Running (8 mph)	13.76	6.26	1121

4.7 Nuclear Energy

We are all very comfortable with the concept of gravity. Using some calculus and the simple notion that the gravitational attraction between two objects is proportional to the product of their masses and inversely proportional to the square of their separation, Newton and Kepler were able to predict the behavior of the solar system. Similarly, in the electrical world, we know opposite charges attract with a force proportional to their charges and inversely proportional to the distance between them. Both gravitational and electrical forces exist within a nucleus, so we might wonder how these forces play out against each other. The gravitational attraction between nucleons pulls them together while the electrical forces between like-charged particles try to shove protons apart. The gravitational attraction is far weaker than the electrical repulsion, which suggests the nucleus should fly apart. But, of course, it doesn't.

There are, it turns out, extremely powerful forces that hold a nucleus together. The energy that would be required to break apart that nucleus, separating it into individual protons and neutrons, is called the nuclear *binding energy*. And this is where Einstein's famous relationship comes in.

Eq. 4.25

$$E = mc^2$$

where E = energy (kJ)
m = mass (kg)
c = the speed of light = 2.998×10^8 m/s

If we imagine building an atom out of individual electrons, protons, and neutrons we would find out that the resulting atom has less mass than the sum of the masses of its individual constituents. That difference in mass, converted to energy units using Equation 4.25 is the binding energy of the atom.

4.7.1 The Nature of Radioactivity

For many people, the mere mention of radioactivity conjures up images of cancer, birth defects, and mushroom clouds. But radioactivity also benefits humankind. Radioactive elements (called *radioisotopes* or *radionuclides*) are used as labels or tags to help unravel the complexities of chemical reactions; indeed, such use was crucial to developing our present understanding of the utterly important process of photosynthesis. Radioisotopes also provide an accurate way to date past events, both historical and geological. And, radioisotopes offer an important way to focus radiation directly onto tumors to help cure those afflicted with cancer.

The term *radioactivity* refers to an atomic nucleus that is unstable. In trying to reach a more stable configuration of protons and neutrons in its nucleus, a radioactive atom emits various forms of radiation, transforming itself from one chemical element into another. The first experiments demonstrating the spontaneous decay of a naturally occurring element were reported by Henri Becquerel in 1896 using a uranium-containing ore called pitchblende. The radiation that Becquerel observed was the emission of what is now called an *alpha* (α) particle. An alpha particle is a two-proton, two-neutron helium nucleus, and the reaction he observed can be described as follows:

Eq. 4.26 $$^{238}_{92}\mathrm{U} \rightarrow {}^{234}_{90}\mathrm{Th} + {}^{4}_{2}\mathrm{He}\ (\alpha\ \text{particle}) \qquad T_{1/2} = 4.51 \times 10^{9}\ \text{yrs}$$

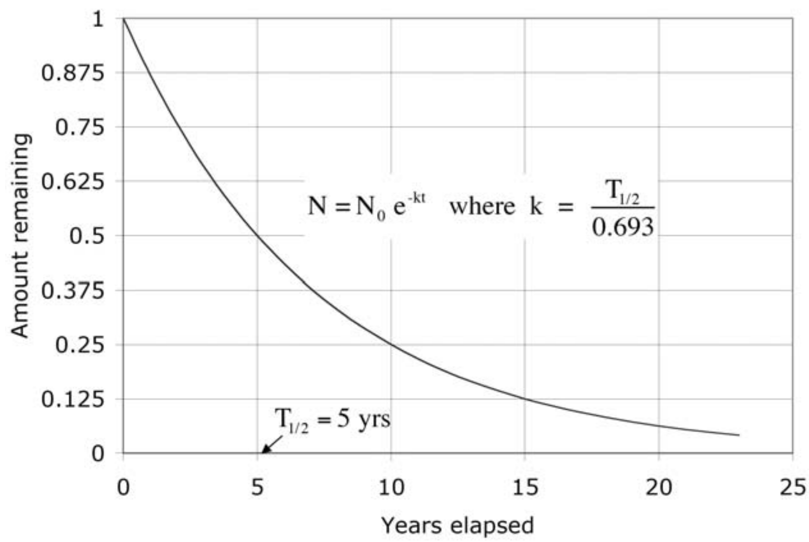
Recall that the number to the lower-left of a chemical symbol is the number of protons in the nucleus whereas that to the upper-left is the sum of protons plus neutrons. When a U-238 atom (with 92 protons) emits an alpha particle, it loses two protons so it becomes a new element—thorium, which has 90 protons. And because the nucleus has lost four nucleons, its mass number drops from 238 to 234.

Also shown in Equation 4.26 is the *half-life* of U-238, which is about 4.5 billion years. That is, if we had 1 kg of U-238 today, in 4.5 billion years half of that will be transformed into thorium and we would have left only 0.5 kg of uranium. After 9 billion years, we would have half of that left, or 0.25 kg. Figure 4.9 illustrates the concept of half-life.

As an alpha particle passes through an object, its energy is gradually dissipated as it interacts with other atoms. Its positive charge attracts electrons in its path, raising their energy levels and possibly removing them completely from their nuclei (*ionization*). Alpha particles are relatively massive and easy to stop. Our skin is sufficient protection for sources that are external to the body, but taken internally, such as by inhalation, alpha particles can be extremely dangerous.

A radionuclide can decay in other ways besides the emission of an alpha particle. Many radionuclides have too many neutrons relative to their number of protons and decay by converting one of those neutrons into a proton plus an electron, with the electron being emitted from the nucleus. These negatively charged electrons are referred to as *negative beta* (β^-) particles. Emission of a β^- results in an increase in the atomic number by one, whereas

figure 4.9



The half-life of a radioactive isotope ($T_{1/2}$) is the time required for half of the material to be converted to another element. As shown, a material with a half-life of 5 years would have half remaining after 5 years, one-fourth after 10 years, one-eighth after 15 years, and so forth.

the mass number remains unchanged. The following reaction shows the decay of strontium-90 into yttrium-90:

Eq. 4.27 $${}^{90}_{38}Sr \rightarrow {}^{90}_{39}Y + \beta^- \qquad T_{1/2} = 29 \text{ yrs}$$

It is also possible for a nucleus to have too many protons, in which case a proton may change into a neutron while ejecting a positively charged electron, or positron, designated as β^+. An example of this reaction is the conversion of nitrogen-13 into carbon-13:

Eq. 4.28 $${}^{13}_{7}N \rightarrow {}^{13}_{6}C + \beta^+ \qquad T_{1/2} = 9.96 \text{ min}$$

As β particles pass through materials, they are also capable of ionizing atoms in our tissues, and they may do so at much greater depths. While alpha particles may travel less than 100 μm into tissue, β radiation may travel several centimeters. They can be stopped with a modest amount of shielding, however. For example, a centimeter or so of aluminum is sufficient.

Equations 4.26–4.28 illustrate the spontaneous emission of actual particles having mass; usually there will also be electromagnetic *gamma* (γ) radiation released. Gamma rays have very short wavelengths in the range of 10^{-11} to 10^{-13} m. Having such short wavelengths means that individual photons are highly energetic and easily cause biologically damaging ionizations. These rays are difficult to contain and may require several centimeters of lead to provide adequate shielding.

All of these forms of ionizing radiation are dangerous to living things. The electron excitations and ionizations that are caused by such radiation cause molecules to become unstable, resulting in the breakage of chemical bonds and other molecular damage. The chain of chemical reactions that follows creates new molecules that did not exist before the irradiation. Exposure to ionizing radiation can result in everything from cancer, leukemia, sterility, cataracts, and reduced lifespan to mutations in chromosomes and genes that will be transmitted to future generations.

4.7.2 Nuclear Fission

All elements having more than 83 protons are naturally radioactive, spitting out combinations of alpha, beta, and gamma radiation. In addition, other nuclear reactions offer us the tantalizing potential to tap into the energy within the nucleus itself. Just before World War II, nuclear scientists discovered that one particular isotope of uranium, U-235, can *fission*, or break apart, when bombarded with neutrons. As suggested in Figure 4.10, when U-235 absorbs a neutron it becomes unstable U-236, which almost instantly splits apart, discharging two *fission fragments* along with two or three neutrons and an intense burst of gamma (γ) rays. Most of the energy released is in the form of kinetic energy in the fission fragments. In a nuclear reactor, that kinetic energy is used to boil water to make steam to spin a turbine and generator.

The fission fragments produced are always radioactive, and concerns for their safe disposal have created much of the controversy surrounding nuclear reactors. Typical fission fragments include cesium-137, which concentrates in muscles and has a half-life of 30 years, and strontium-90, which concentrates in bone and has a half-life of 8.1 days. The half-lives of fission fragments tend to be no longer than a few decades, so after a period of several hundred years their radioactivity will decline to relatively insignificant levels.

An example of one such fission is the following reaction:

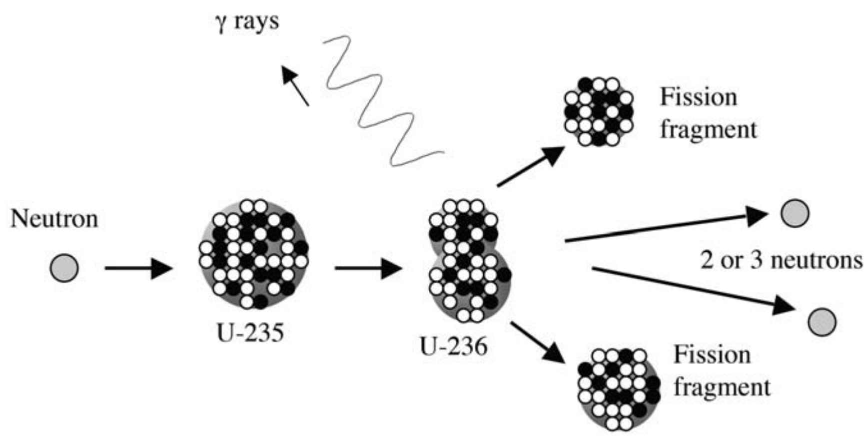
Eq. 4.29

$$^{235}_{92}\mathrm{U} + \mathrm{n} \rightarrow \,^{143}_{55}\mathrm{Cs} + \,^{90}_{37}\mathrm{Rb} + 3\mathrm{n} + 184\ \mathrm{MeV}$$

Notice that there is a balance of nucleons on each side of the reaction and there is an amount of energy released, here expressed in millions of electron-volts (MeV). Although Equation 4.20 expresses the conversion of mass to energy using units of kJ, that unit is hardly ever used in calculations associated with nuclear reactions because it is just too big to be convenient. The more common measure is *electron-volts* (eV), or million electron-volts (MeV), where $1\ \mathrm{eV} = 1.60 \times 10^{-19}\ \mathrm{J}$. The basis for the electron-volt measurement will be described in section 4.8, "Electrical Energy" and an example of its use is demonstrated in Solution Box 4.9.

Immediately apparent from Figure 4.10 is that it took a single neutron to cause one atom of U-235 to fission, and in doing so two or three new neutrons are emitted. Those neutrons can go on to cause other fissions, leading to the possibility of a controlled, self-

figure 4.10



The fissioning of U-235 creates two radioactive fission fragments plus two or three neutrons and gamma rays.

sustaining chain reaction for a nuclear power plant, or an uncontrolled, explosive chain reaction as occurs in an atomic bomb. To create a uranium bomb, however, requires a much more concentrated supply of U-235 than is present in a conventional nuclear power plant.

Uranium-235 is the only fissile material occurring in nature, but it makes up only about 0.7% of naturally occurring uranium. The other 99.3% is mostly U-238, which does not fission. To build a nuclear reactor requires a U-235 concentration of about 3%, but to build a bomb it needs to be enriched to more than 90%. That enrichment process is extremely difficult and is a major deterrent to creating bomb-grade material from U-235. There is, however, another approach that is based on creating fissile plutonium in a conventional reactor and then separating it out from the reactor wastes. Indeed, this is how the bomb that destroyed Nagasaki was built, and it is the process used by every country that now possesses nuclear weapons.

Most of the uranium in a reactor is the isotope U-238, but it is not capable of fissioning. However, when it absorbs a neutron, it can be transformed through the following reactions to plutonium, and plutonium is a fissile material (Figure 4.11).

Eq. 4.30

$$^{238}_{92}\mathrm{U} + \mathrm{n} \rightarrow {}^{239}_{92}\mathrm{U} \xrightarrow{\beta} {}^{239}_{93}\mathrm{Np} \xrightarrow{\beta} {}^{239}_{94}\mathrm{Pu}$$

Plutonium is radioactive, emits alpha particles when it decays, and has a half-life of 24,360 years. It does not occur in nature and the plutonium we produce will be around for many tens of thousands of years. It is a worrisome material not only because it can be used for nuclear weapons, but also because it is considered by many to be the most toxic substance known to humankind. Its presence in nuclear wastes, along with other long-lived actinides, contributes to the need to isolate those wastes, in essence, forever.

By converting exceedingly small amounts of mass into exceptionally large amounts of energy, nuclear energy gives us the potential to wean ourselves from our fossil fuel

SOLUTION BOX 4.9

Nuclear Energy versus Chemical Energy

A typical fission reaction releases about 200 MeV per atom of U-235. How does this compare with the energy released by burning methane?

Solution:

For the U-235, first convert that 200 MeV per atom to joules:

$$\text{U-235: } 200\text{ MeV} \times 1.60 \times 10^{-22}\text{ kJ/eV} \times 10^{6}\text{ eV/MeV} = 3.2 \times 10^{-14}\text{ kJ/atom}$$

Those uranium atoms are very heavy, so converting to a per unit of mass basis

$$\frac{3.2 \times 10^{-14}\text{ kJ/atom} \times 6.023 \times 10^{23}\text{ atoms/mol}}{235\text{ g/mol}} = 82.02 \times 10^{6}\text{ kJ/g U-235}$$

Recall from Solution Box 4.5 that burning a mole of CH_4 liberates 802 kJ, which translates to

$$CH_4\text{: } \frac{802\text{ kJ/mol}}{(12 + 4 \times 1)\text{ g/mol}} = 50.125\text{ kJ/g } CH_4$$

Wow! The nuclear reaction with U-235 releases more than 1.6 million times as much energy per gram as the chemical reaction with CH_4.

A similar assessment for coal, accounting for the concentration of U-235 in nuclear fuel, and a representative type of coal, works out to about 1 ton of nuclear fuel providing the same energy as roughly 100,000 tons of coal.

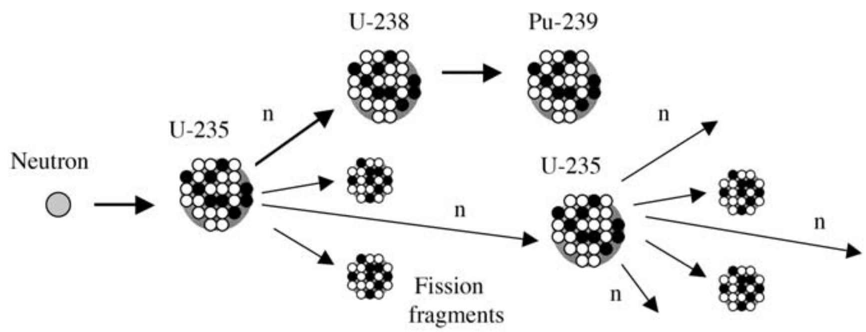
SOLUTION

dependence with all of its environmental and national security disadvantages. As usual, however, we can't get something for nothing and nuclear power presents its own challenges having to do with disposal of radioactive wastes, human and financial risks should there be a reactor accident, and the increasing worry about terrorists gaining access to plutonium to make their own atomic bombs.

4.7.3 Nuclear Fusion

Nuclear fission relies on the ability to break atoms apart; the opposite occurs in nuclear fusion reactions. With fusion, it is the loss of mass associated with merging nuclei that creates the energy

figure 4.11



Plutonium created in a nuclear reactor is a source of fissile material that can be used either as a reactor fuel or for nuclear weapons.

we are after. The sun is the perfect example of a safe, working nuclear fusion device; a not-so-perfect example is a thermonuclear bomb. The reactions occurring within the sun involve a number of steps, but can be summarized as four hydrogen atoms combining to form helium-4:

Eq. 4.31
$$4\,{}^{1}_{1}\mathrm{H} \rightarrow {}^{4}_{2}\mathrm{He} + 2\beta^{+} + \text{energy}$$

The most promising fusion reactions for us down here on Earth involve various isotopes of hydrogen, which combine to form helium. Hydrogen has only one proton, and in nature it almost always has no neutrons. A very small fraction, however, about 0.015% of naturally occurring hydrogen, has a single neutron in the nucleus along with the proton, and this isotope is given a special name, *deuterium*. Another important isotope has 2 neutrons, and it is called *tritium*. A sample deuterium-deuterium (D-D) reaction is the following:

Eq. 4.32
$${}^{2}_{1}\mathrm{H} + {}^{2}_{1}\mathrm{H} \rightarrow {}^{3}_{2}\mathrm{He} + \mathrm{n} + 3.3\ \mathrm{MeV}$$

Because roughly one out of every 5000 hydrogen atoms in water is deuterium, there is enough deuterium in the oceans to supply all of the world's energy needs for millions of years into the future. This deuterium-deuterium (D-D) reaction presents an extreme challenge because the reactions need to take place under conditions that are even hotter than the interior of the sun.

A much more promising first-generation fusion reactor is based on fusing deuterium and tritium in the following D-T reaction:

Eq. 4.33
$${}^{2}_{1}\mathrm{H} + {}^{3}_{1}\mathrm{H} \rightarrow {}^{4}_{2}\mathrm{He} + \mathrm{n} + 17.6\ \mathrm{MeV}$$

The trick in this case is finding enough tritium. Tritium has a half-life of only 12 years and does not exist in any great quantity in nature, so if this reaction is to provide much energy for the future we will have to find some way to manufacture it. One way to produce tritium is to expose lithium-6 to neutron bombardment in a fission reactor, as follows:

Eq. 4.34
$$n + {}^{6}_{3}\mathrm{Li} \rightarrow {}^{3}_{1}\mathrm{H} + {}^{4}_{2}\mathrm{He} + 4\ \mathrm{MeV}$$

Unfortunately, there isn't that much lithium in the world either. The United States, for example, might be able to produce enough lithium from its own resources to power the country for only a few hundred years. That's pretty good, of course, but not as exciting as the dazzling potential of D-D fusion.

Significant progress has been made on D-T fusion, to the point that in mid-2005, it was announced that a consortium made up of the European Union, the United States, Russia, Japan, South Korea, and China had agreed to build a demonstration reactor called the International Thermonuclear Experimental Reactor (ITER), to be constructed in Cadarache, France. It will be something like building our own small star, with reaction temperatures hotter than our own sun, around 100 million kelvins. If all goes well, it will become operational in 2015 at a cost of around $15 billion, and could eventually lead to commercial reactors that would have the potential to supply enormous amounts of energy without greenhouse gases and far less radioactivity than current fission reactors.

4.8 Electrical Energy

At last count, there were only four fundamental forces in the universe. The first is gravitational force. Two others are forces within atoms, called the strong nuclear force and the weak nuclear force. And the fourth is the electrical force that one charged object exerts on another. That fourth force was first explained by the French physicist Charles-Augustin de Coulomb (1736–1806) and the result is known as *Coulomb's law:*

Eq. 4.35
$$F = k\frac{q_1 q_2}{d^2}$$

where, in SI units, F is in newtons (N), q_1 and q_2 are the electrical charges of each object, and are given in coulombs (C). The distance d between the two objects is in meters, and the coefficient k is 9×10^9 (N-m^2/C^2).

To visualize this phenomenon, it helps to imagine an invisible *electric field* surrounding every charge. When a test charge is brought into an electric field we can measure the force exerted by the field on the charge, so we know it is there even if we can't see it. This simple notion that electric fields exist and they can exert forces on charges has implications that extend throughout the field of electrical engineering.

4.8.1 Electric Current

One coulomb of charge is equal to the charge on 6.242×10^{18} electrons—that's a lot of electrons—but, when it comes to having electrons do work for us, some are much more helpful than others. Most electrons are too tightly bound to their nuclei to do us much good, but others are far enough away that the attraction of any particular nucleus can easily be overcome. These *free electrons* easily wander from one atom to another, and if subject to even a slight electric field they can be made to bounce along in a particular direction. That flow of charge constitutes an electric current.

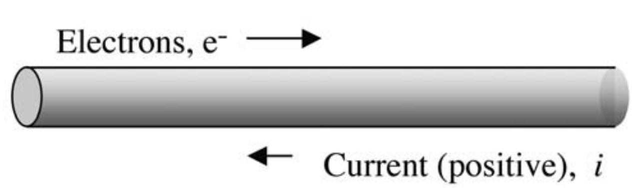
In general, charges can be negative or positive. In a copper wire, the only charge carriers are electrons, with negative charge. In a neon light, however, under the influence of an electric field positive ions move in one direction and negative electrons move in the other. Each contributes to current and the total current is their sum. By convention, the direction of current flow is taken to be the direction that positive charges would flow, whether or not positive charges happen to be in the picture. Thus, in a wire, electrons moving in one direction constitute a current flowing in the opposite direction, as shown in Figure 4.12.

If we imagine a wire, when one coulomb's worth of charge passes a given spot in one second the current is defined to be one *ampere* (abbreviated A), named after the nineteenth-century physicist Andre Marie Ampere. In equation form, current i is the flow of charge q past a point, or through an area, per unit of time t.

Eq. 4.36
$$i = \frac{q}{t}$$

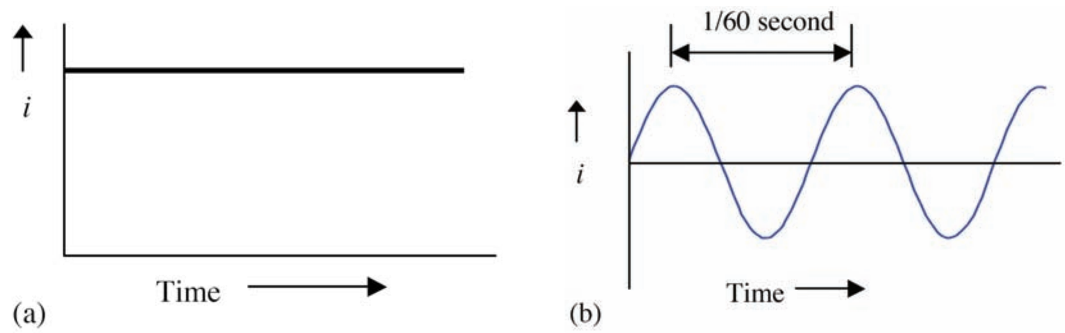
When charge flows at a steady rate in one direction only, it is said to be *direct current*, or *dc*. A battery, for example, supplies direct current. When charge flows back and forth sinusoidally, it is called *alternating current*, or *ac*. In the United States the ac electricity delivered by the power company has a frequency of 60 cycles per second, or 60 hertz (abbreviated Hz). Examples of dc and ac are shown in Figure 4.13.

figure 4.12



By convention, in a wire the direction of positive current flow is the opposite of the actual flow of electrons.

figure 4.13 **Direct Current and Alternating Current**



(a) Steady, direct current (dc). (b) 60 Hz alternating current (ac).

4.8.2 Voltage

Electrons won't flow through a circuit unless they are given some energy to help them along their way. That "push" is measured in volts, where voltage is defined to be the amount of energy (w, joules) given to a unit of charge, q.

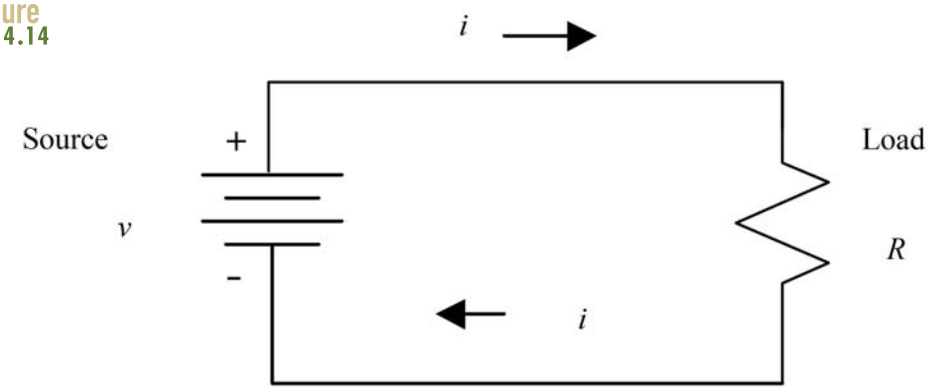
Eq. 4.37
$$v = \frac{w}{q}$$

For example, a 9-volt battery provides 9 joules of energy to each coulomb of charge that it stores. Voltage describes the potential for charge to do work. Just as mechanical forms of potential energy are always measured with respect to some reference, so too is voltage. Thus, the positive terminal of a 9-volt battery is 9 volts higher than the voltage of the negative terminal. The negative terminal may be grounded by connecting a wire from it to a copper rod pounded into the Earth, in which case we would say the positive terminal is 9 volts with respect to ground.

4.8.3 The Concept of an Electrical Circuit

A simple circuit consists of a *source* of energy, say a battery; a *load* such as a lightbulb, toaster, or something else that you want to power with electricity; and some connecting wire to carry current from the source to the load and back again. Notice that a circuit has to have a return path for the electrons. They can't just go one way to the lightbulb and then drop off onto the floor.

figure 4.14



A simple dc circuit consisting of a voltage source, a resistive load, and a connecting wire through which current flows.

In a simple dc circuit, such as the one shown in Figure 4.14, the load can be characterized by its *resistance* to the flow of electrons through it. The wires, on the other hand, are often assumed to be perfect conductors that offer no resistance at all to electron flow. The relationship between the voltage v applied to this circuit, the current that flows through the circuit i, and the resistance offered by the load to that current flow R (ohms, Ω) is given by

Eq. 4.38 $$v \text{ (volts)} = R \text{ (ohms)} \times i \text{ (amps)}$$

This deceptively simple relationship is known as Ohm's law in honor of the German physicist, Georg Ohm, whose original experiments led to this incredibly important relationship.

4.8.4 Electrical Power and Energy

Electric circuits can have many different goals, but they can be distinguished by whether their purpose is to transmit information or power. Most of the circuitry in your laptop is there to process information by detecting the presence or absence of a voltage (0s and 1s). In that context, power consumption is a bad thing because it makes your battery drain faster. In the context of this book, however, our goal is to generate and transmit large amounts of electrical energy and power to do real work, work that can not only charge those laptop batteries, but also power whole factories, cities, and regions of the country.

If we go back to the simple circuit of Figure 4.14, we can derive a few important relationships between voltage, resistance, and current and then relate those to power and energy. Recall that voltage is the energy given to a unit of charge ($v = w/q$) and current is the rate at which charge moves around a circuit (q/t). Combining these two with the definition of power, which is the rate of delivering energy ($p = w/t$), we get:

Eq. 4.39

$$p = \frac{w}{t} = \frac{w}{q} \cdot \frac{q}{t} = vi$$

So, the power delivered to a load is the product of the voltage across the load times the current through the load. Combining Equation 4.39 with Ohm's law (4.38) gives us three ways to express power to a resistive load:

Eq. 4.40

$$p = vi = \frac{v^2}{R} = i^2 R$$

When v is in volts, i is in amps, and R is in ohms, the above expressions give power in watts (1 W = 1 J/s).

Multiplying power watts (W) or kilowatts (kW) by the length of time that the power is consumed (hours) we get energy in watt-hours (Wh) or kilowatt-hours (kWh). The example in Solution Box 4.10 illustrates the above electrical relationships.

In Chapter 11, we'll be looking at how much electrical power and energy various household appliances use and then figuring out how much that will cost. The process is conceptually simple. Just multiply the watts used by an appliance by the hours of operation and you have watt-hours. Household utility bills are based on how many kilowatt-hours you use, so divide that by 1000 to get kWh.

As the example in Solution Box 4.11 illustrates, some appliances use power whether they are turned on or not. Look around your home and see how many devices have glowing red or green lights turned on all the time. Your TV and other remote controlled appliances consume standby power while their electronics wait for you to push the "on" button from your remote. In fact, upwards of two-thirds of the energy used in many home electronic devices is used for standby power.

4.9 Summary

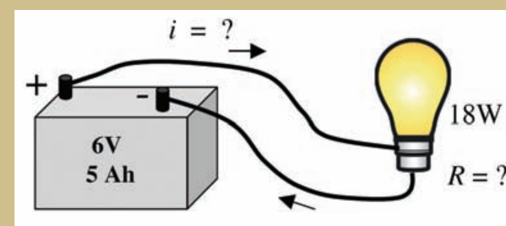
In this chapter on fundamentals, we have tried to lay the groundwork you need to understand the wide range of energy technologies that will follow. You should now understand energy units, conversions between them, and ways to describe very small values (e.g., nanometers) and very large quantities (e.g., exajoules). We also made an important distinction between energy and power. Power is a rate; that is, it is energy per unit of time. Some power units sound like rates (e.g., Btu per hour) whereas others do not (e.g., kilowatts). You can lose credibility if you use them wrongly (e.g., kilowatts per hour).

We have explored the ways that nature manages to store energy in the potential and kinetic forms associated with work, in heat energy that is transferred from warm objects

SOLUTION BOX 4.10

Power to an Incandescent Lamp

Suppose a heavy-duty flashlight uses a 6-volt, lead-acid battery to deliver power to the filament of an 18-watt bulb. Find the resistance of the filament and the current that flows.



Solution:

Because power is volts times amps, the current drawn by the bulb will be

$$i = \frac{18\ \mathrm{W}}{6\ \mathrm{V}} = 3\ \mathrm{A}$$

From Ohm's law, $v = R\,i$, the filament resistance must be

$$R = \frac{v}{i} = \frac{6\ \mathrm{V}}{3\ \mathrm{A}} = 2\ \Omega$$

Batteries are usually rated by their amp-hour (Ah) capacity, which roughly speaking translates to the product of the amps they deliver multiplied by the hours they can provide those amps. If the battery in this flashlight is rated at 5 Ah, how long could it keep the flashlight going? How much energy would have been delivered?

$$\text{Time} = \frac{5\ \mathrm{Ah}}{3\ \mathrm{A}} = 1.67\ \mathrm{hr}$$

$$\text{Energy} = 18\ \mathrm{W} \times 1.67\ \mathrm{hr} = 30\ \mathrm{Wh}$$

to cold ones, in the chemical form stored in the bonds between atoms, in the electromagnetic form from the sun that powers essentially all living things, in the form released when fission breaks apart atoms or fusion joins them together, and finally in the electricity that transformed the twentieth century, enabling all of the technology marvels we so enjoy today.

In nature, and in physical systems that humans construct, energy is constantly being shuttled about from one form to another. The first law of thermodynamics tells us that energy is conserved; that is, we can draw up balance sheets that account for every bit of it, including the conversion of mass to energy. To say we have consumed, or "used up," some energy is therefore a bit of a misnomer. We merely transform it, but while we do so we are constantly degrading the quality of energy, which is where the second law of thermodynamics comes in.

SOLUTION

SOLUTION BOX 4.11

That Satellite TV System

A typical satellite TV system consumes 16 watts while turned off and 17 watts when it is in use. Similarly, a typical 20-inch TV consumes 5 watts while off and 68 watts when on.

Suppose you watch TV 4 hours per day, 30 days per month. What fraction of your electricity is used while the TV is turned off?

16 W off, 17 W on   5 W off, 68 W on

Solution:

TV on: **(68 W + 17 W) × 4 h/d × 30 d/mo = 10,200 Wh/mo = 10.2 kWh/mo**

TV off: **(5 W + 16 W) × 20 h/d × 30 d/mo = 12,600 Wh/mo = 12.6 kWh/mo**

Total = 22.8 kWh/mo

So, **12.6/22.8 = 0.553 = 55.3%** of your electricity is used while the TV is off.

If the utility charges 10¢ per kWh, how much does your TV cost you in a 30-day month? What is that cost per hour of actual TV use?

At **$0.10/kWh × 22.8 kWh/mo = $2.28/mo**

Per hour of use that works out to **228¢/120 hr = 1.9¢/hour**

The first law deals with energy quantities and the second with energy quality. As we manipulate energy, the quality of that energy is constantly degraded; that is, more and more of it ends up as relatively useless waste heat. The second law also provides information on the direction in which energy flows may take place. On your kitchen table, it has no problem with hot coffee cooling off and cold beer warming up, but it frowns on allowing the opposite to spontaneously occur. A more subtle interpretation of the second law describes the universe as constantly moving toward greater and greater disorder. The orderliness of hot separated from cold, becomes the random sameness of warm. The orderly arrangement of molecules in ice becomes a collection of disordered, randomly arranged molecules when ice melts.

In subsequent chapters, the energy flow concepts introduced here will be applied to real systems. Questions such as how we size systems, estimate their performance, evaluate their environmental implications, and determine their economic value will be addressed.