11 Using Energy

The odd hopping gait of a kangaroo has a very practical purpose: It lets the kangaroo cover great distances with minimal energy input. How is this efficiency possible?

LOOKING AHEAD ▸▸

The goals of Chapter 11 are to learn more about energy transformations and transfers, the laws of thermodynamics, and theoretical and practical limitations on energy use.

Efficiency

In this chapter, we'll look at practical limits on energy transfers and transformations.

Looking Back ◂◂
10.1 Basic energy concepts; forms of energy
10.6 The law of conservation of energy

Both bulbs put out the same amount of light, but the one on the right uses 1/4 the electric power. Both bulbs perform the same transformation, but one is much more efficient.

Energy in the Body

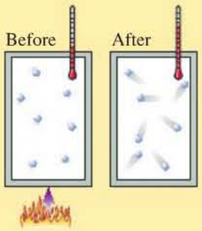
All the energy that your body uses for all of the tasks you complete during the day comes from food. How efficient is your body at converting this energy? How much energy does your body actually use to run, climb, and move?

Climbing stairs requires a change in potential energy. How much energy does your body use to make this climb?

These fighters wear masks that allow a direct measurement of the energy used by their bodies as they exercise.

Thermal Energy and Temperature

An object's temperature is related to its thermal energy, the energy of motion on an atomic scale.

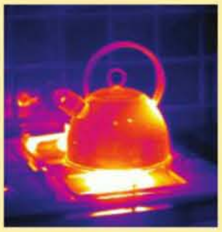


We'll use the ideal gas model to help us understand the nature of thermal energy and temperature.

Looking Back ◂◂
10.7 Thermal energy

Heat and Thermodynamics

Processes in which only thermal energy changes are the domain of **thermodynamics.**



The thermal energy of the kettle is increased by the heat from the burner, which is at a higher temperature.

Heat Engines and Heat Pumps

A **heat engine** can convert thermal energy into other forms of energy. A **heat pump** moves thermal energy from one place to another.

This geothermal plant uses volcanic thermal energy to generate electricity. "Waste" heat warms the lagoon. Why must any energy be "wasted"?

The inside of a refrigerator is cold because heat has been "pumped" out. What's the energy cost to move this heat?

Entropy

Entropy is a measure of disorder at an atomic level that helps us explain some basic observations about the world.

When you stir a cup of coffee, the cream mixes in. This is irreversible; stirring backward won't cause it to unmix. The concept of entropy explains why the future is different from the past and why there are theoretical limits on energy use.

11.1 Transforming Energy

As we saw in Chapter 10, energy can't be created or destroyed; it can only be converted from one form to another. When we say we are *using energy,* we mean that we are transforming it, such as transforming the chemical energy of food into the kinetic energy of your body. Let's revisit the idea of energy transformations, considering some realistic situations that have interesting theoretical and practical limitations.

Energy transformations

Light energy hitting a solar cell on top of these walkway lights is converted to electric energy and then stored as chemical energy in a battery.

Light energy is absorbed by photosynthetic pigments in soybean plants, which use this energy to create concentrated chemical energy.

A wind turbine converts the translational kinetic energy of moving air into electric energy.

At night, the battery's chemical energy is converted to electric energy that is then converted to light energy in a light-emitting diode.

The soybeans are harvested and their oil is used to make candles. When the candle burns, the stored chemical energy is transformed into light energy and thermal energy.

Miles away, this energy can be used by a fan, which transforms the electric energy back into the kinetic energy of moving air.

In each of the processes in the table, energy is transformed from one form into another, ending up in the same form as it began. But some energy appears to have been "lost" along the way. The garden lights certainly shine with a glow that is much less bright than the sun that shined on them. No energy is really lost because energy is conserved; it is merely converted to other forms that are less useful to us.

The table also illustrates another key point: We are broadening our scope to include forms of energy beyond those we considered in Chapter 10, including radiant energy (the light that hits the solar cells and the plants) and electric energy (energy that is transferred from the windmill to the fan).

Recall the work-energy equation from Chapter 10:

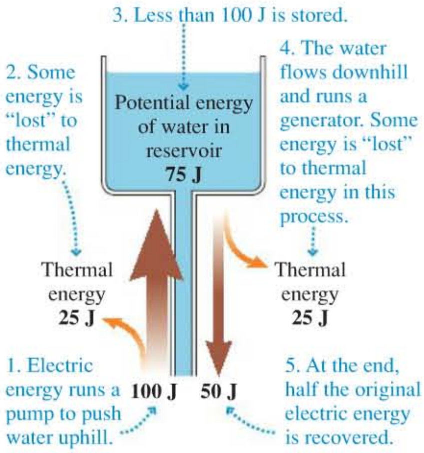
$$\Delta E = \Delta K + \Delta U + \Delta E_{\text{th}} + \Delta E_{\text{chem}} + \cdots = W \qquad (11.1)$$

This equation includes work, an energy transfer. Work can be positive or negative; it's positive when energy is transferred into a system, negative when energy is transferred out of a system. In Chapter 10 we defined work as a mechanical transfer of energy; in this chapter we'll broaden our definition of work to include electric energy into and out of motors and generators.

Electric power companies sometimes use excess electric energy to pump water uphill to a reservoir. They reclaim some of this energy when demand increases by letting the water generate electricity as it flows back down. **FIGURE 11.1** shows the energy transformations in this process, assuming 100 J of electric energy at the start. We define the system as the water and the network of pipes, pumps, and generators that makes up the plant. The electric energy to run the pump is a work input, and so it is positive; the electric energy that is extracted at the end is a work output, and so is negative. Pumping the water uphill requires 100 J of work; this increases the potential energy of the water by only 75 J. The "lost" energy is simply the energy transformed to thermal energy due to friction in the pump and other causes, as we can see by applying Equation 11.1 to this situation:

$$\Delta E_{\text{th}} = W - \Delta U = 100\text{ J} - 75\text{ J} = 25\text{ J}$$

FIGURE 11.1 Energy transformations for a pumped storage system.



When the water flows back downhill, the potential energy decreases, meaning a negative value of ΔU. Energy is transferred out of the system as work, meaning a negative value for W in the conservation of energy equation as well. The change in thermal energy in this case is

$$E_{th} = W - \Delta U = (-50 \text{ J}) - (-75 \text{ J}) = 25 \text{ J}$$

For the pumped storage scheme of Figure 11.1, as for the cases in the table on the previous page, the energy output is in the same form as the energy input, but some energy appears to have been "lost" as the transfers and transformations took place. The power company put in 100 J of electric energy at the start, but only 50 J of electric energy was recovered at the end. No energy was actually lost, of course, but, as we'll see, when other forms of energy are transformed into thermal energy, this change is *irreversible*: We cannot easily recover this thermal energy and convert it back to electric energy. In other words, **the energy isn't lost, but it is lost to our use.** In order to look at losses in energy transformations in more detail, we will define the notion of efficiency.

Efficiency

To get 50 J of energy out of your pumped storage plant, you would actually need to *pay* to put in 100 J of energy, as we saw in Figure 11.1. Because only 50% of the energy is returned as useful energy, we can say that this plant has an efficiency of 50%. Generally, we can define efficiency as

$$e = \frac{\text{what you get}}{\text{what you had to pay}} \tag{11.2}$$

General definition of efficiency

The larger the energy losses in a system, the lower its efficiency.

Reductions in efficiency can arise from two different sources:

- **Process limitations.** In some cases, energy losses may be due to practical details of an energy transformation process. You could, in principle, design a process that entailed smaller losses.
- **Fundamental limitations.** In other cases, energy losses are due to physical laws that cannot be circumvented. The best process that could theoretically be designed will have less than 100% efficiency. As we will see, these limitations result from the difficulty of transforming thermal energy into other forms of energy.

If you are of average weight, walking up a typical flight of stairs increases your body's potential energy by about 1800 J. But a measurement of the energy used by your body to climb the stairs would show that your body uses about 7200 J to complete this task. The 1800 J increase in potential energy is "what you get"; the 7200 J your body uses is "what you had to pay." We can use Equation 11.2 to compute the efficiency for this action:

$$e = \frac{1800 \text{ J}}{7200 \text{ J}} = 0.25 = 25\%$$

This relatively low efficiency is due to *process limitations.* The efficiency is less than 100% because of the biochemistry of how your food is digested and the biomechanics of how you move. The process could be made more efficient by, for example, changing the angle of the stairs.

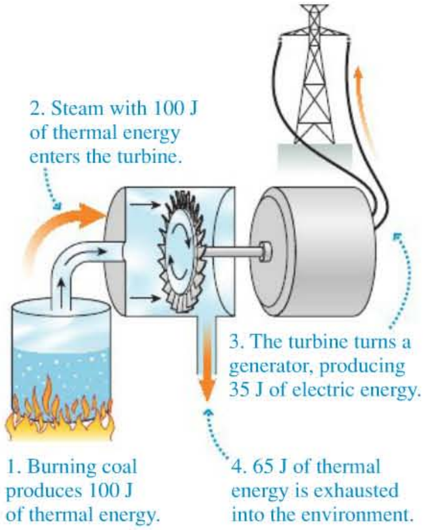
These cooling towers release thermal energy from a coal-fired power plant.

The electric energy you use daily must be generated from other sources—in most cases the chemical energy in coal or other fossil fuels. In a coal-fired power plant, the chemical energy is converted to thermal energy by burning, and the resulting thermal energy is converted to electric energy. A typical power-plant cycle is shown in FIGURE 11.2. "What you get" is the energy output, the 35 J of electric energy. "What you had to pay" is the energy input, the 100 J of chemical energy, giving an efficiency of

$$e = \frac{35\ \text{J}}{100\ \text{J}} = 0.35 = 35\%$$

Unlike your stair-climbing efficiency, the rather modest power-plant efficiency turns out to be largely due to a *fundamental limitation:* Thermal energy cannot be transformed into other forms of energy with 100% efficiency. A 35% efficiency is close to the theoretical maximum, and no power plant could be designed that would do better than this maximum. In later sections, we will explore the fundamental properties of thermal energy that make this so.

FIGURE 11.2 Energy transformations in a coal-fired power plant.



PROBLEM-SOLVING STRATEGY 11.1 **Energy efficiency problems**

PREPARE There are two key components to define before we compute efficiency:

1. Choose what energy to count as "what you get." This could be the useful energy output of an engine or process or the work that is done in completing a process. For example, when you climb a flight of stairs, "what you get" is your change in potential energy.
2. "What you had to pay" will generally be the total energy input needed for an engine, task, or process. For example, when you run your air conditioner, "what you had to pay" is the electric energy input.

SOLVE You may need to do additional calculations:

- Compute values for "what you get" and "what you had to pay."
- Be certain that all energy values are in the same units.

After this, compute the efficiency using $e = \dfrac{\text{what you get}}{\text{what you had to pay}}$.

ASSESS Check your answer to see if it is reasonable, given what you know about typical efficiencies for the process under consideration.

EXAMPLE 11.1 **Lightbulb efficiency**

A 15 W compact fluorescent bulb and a 75 W incandescent bulb each produce 3.0 W of visible-light energy. What are the efficiencies of these two types of bulbs for converting electric energy into light?

PREPARE The problem statement doesn't give us values for energy; we are given values for power. But 15 W is 15 J/s, so we will consider the value for the power to be the energy in 1 second.

FIGURE 11.3 Incandescent and compact fluorescent bulbs.

For each of the bulbs, "what you get" is the visible-light output—3.0 J of light every second for each bulb. "What you had to pay" is the electric energy to run the bulb. This is how the bulbs are rated. A bulb labeled "15 W" uses 15 J of electric energy each second. A 75 W bulb uses 75 J each second.

SOLVE The efficiencies of the two bulbs are computed using the energy in 1 second:

$$e(\text{compact fluorescent bulb}) = \frac{3.0\ \text{J}}{15\ \text{J}} = 0.20 = 20\%$$

$$e(\text{incandescent bulb}) = \frac{3.0\ \text{J}}{75\ \text{J}} = 0.040 = 4\%$$

ASSESS Both bulbs produce the same visible-light output, but the compact fluorescent bulb does so with a significantly lower energy input, so it is more efficient. Compact fluorescent bulbs are more efficient than incandescent bulbs, but their efficiency is still relatively low—only 20%.

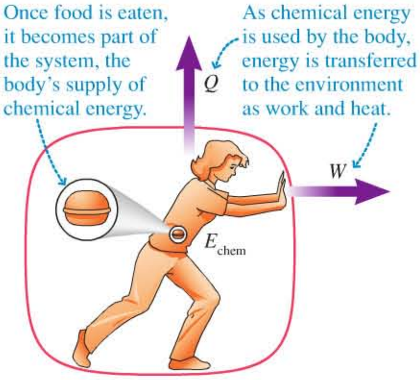
STOP TO THINK 11.1 Crane 1 uses 10 kJ of energy to lift a 50 kg box to the roof of a building. Crane 2 uses 20 kJ to lift a 100 kg box the same distance. Which crane is more efficient?

A. Crane 1.
B. Crane 2.
C. Both cranes have the same efficiency.

11.2 Energy in the Body: Energy Inputs

In this section, we will look at energy in the body, which will give us the opportunity to explore a number of different energy transformations and transfers in a practical context. **FIGURE 11.4** shows the body considered as the system for energy analysis. The chemical energy in food provides the necessary energy input for your body to function. It is this energy that is used for energy transfers with the environment.

FIGURE 11.4 Energy of the body, considered as the system.



Getting Energy from Food

When you walk up a flight of stairs, where does the energy come from to increase your body's potential energy? At some point, the energy came from the food you ate, but what were the intermediate steps? The chemical energy in food is made available to the cells in the body by a two-step process. First, the digestive system breaks down food into simpler molecules such as glucose, a simple sugar, or long chains of glucose molecules called glycogen. These molecules are delivered via the bloodstream to cells in the body, where they are metabolized by combining with oxygen, as in Equation 11.3:

Glucose from the digestion of food combines with oxygen that is breathed in to produce … … carbon dioxide, which is exhaled; water, which can be used by the body; and energy.

$$\underbrace{C_6H_{12}O_6}_{\text{Glucose}} + \underbrace{6O_2}_{\text{Oxygen}} \longrightarrow \underbrace{6CO_2}_{\text{Carbon dioxide}} + \underbrace{6H_2O}_{\text{Water}} + \text{energy} \tag{11.3}$$

This metabolism releases energy, much of which is stored in a molecule called ATP, or adenosine triphosphate. Cells in the body use this ATP to do all the work of life: Muscle cells use it to contract, nerve cells use it to produce electrical signals, and so on.

Oxidation reactions like those in Equation 11.3 "burn" the fuel that you obtain by eating. The oxidation of 1 g of glucose (or any other carbohydrate) will release approximately 17 kJ of energy. Table 11.1 compares the energy content of carbohydrates and other foods to other familiar sources of chemical energy.

TABLE 11.1 Energy in fuels

Fuel	Energy in 1 g of fuel (in kJ)
Hydrogen	121
Gasoline	44
Fat (in food)	38
Coal	27
Carbohydrates (in food)	17
Wood chips	15

It is possible to measure the chemical energy content of food by burning it. Burning food may seem quite different from metabolizing it, but if glucose is burned, the chemical formula for the reaction is Equation 11.3; the two reactions are the same. Burning food transforms all of its chemical energy into thermal energy, which can be easily measured. Thermal energy is often measured in units of **calories (cal)** rather than in joules; 1.00 calorie is equivalent to 4.19 joules.

NOTE ▶ The energy content of food is usually given in Calories (Cal) (with a capital "C"); one Calorie (also called a "food calorie") is equal to 1000 calories or 1 kcal. If a candy bar contains 230 Cal, this means that, if burned, it would produce 230,000 cal (or 964 kJ) of thermal energy. ◀

EXAMPLE 11.2 Energy in food

A 12 oz can of soda contains approximately 40 g (or a bit less than 1/4 cup) of sugar, a simple carbohydrate. What is the chemical energy in joules? How many Calories is this?

SOLVE 1 g of sugar contains 17 kJ of energy; 40 g contains

$$40 \text{ g} \times \frac{17 \times 10^3 \text{ J}}{1 \text{ g}} = 68 \times 10^4 \text{ J} = 680 \text{ kJ}$$

Converting to Calories, we get

$$680 \text{ kJ} = 6.8 \times 10^5 \text{ J} = (6.8 \times 10^5 \text{ J})\frac{1.00 \text{ cal}}{4.19 \text{ J}}$$
$$= 1.6 \times 10^5 \text{ cal} = 160 \text{ Cal}$$

ASSESS 160 Calories is a typical value for the energy content of a 12 oz can of soda (check the nutrition label on one to see), so this result seems reasonable.

Counting calories BIO Most foods burn quite well, as this photo of corn chips illustrates. You could set food on fire to measure its energy content, but this isn't really necessary. The chemical energies of the basic components of food (carbohydrates, proteins, fats) have been carefully measured—by burning—in a device called a *calorimeter.* Foods are analyzed to determine their composition, and their chemical energy can then be calculated.

The first item on the nutrition label on packaged foods is Calories—a measure of the chemical energy in the food. (In Europe, where SI units are standard, you will find the energy content listed in kJ.) The energy content of some common foods is given in Table 11.2.

TABLE 11.2 Energy content of foods

Food	Energy content in Cal	Energy content in kJ	Food	Energy content in Cal	Energy content in kJ
Carrot (large)	30	125	Slice of pizza	300	1260
Fried egg	100	420	Frozen burrito	350	1470
Apple (large)	125	525	Apple pie slice	400	1680
Beer (can)	150	630	Fast-food meal:		
BBQ chicken wing	180	750	burger, fries,		
Latte (whole milk)	260	1090	drink (large)	1350	5660

11.3 Energy in the Body: Energy Outputs

Your body uses energy in many ways. Even at rest, your body uses energy for tasks such as building and repairing tissue, digesting food, and keeping warm. The numbers of joules used per second (that is, the power in watts) by different tissues in the resting body are listed in Table 11.3.

Your body uses energy at the rate of approximately 100 W when at rest. This energy, from chemical energy in your body's stores, is ultimately converted entirely to thermal energy, which is then transferred as heat to the environment. A hundred people in a lecture hall add thermal energy to the room at a rate of 10,000 W, and the air conditioning must be designed to take account of this.

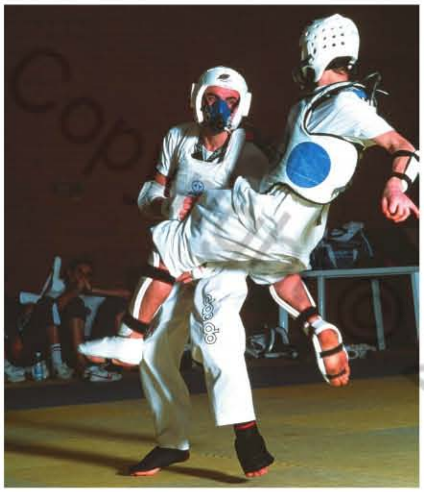
TABLE 11.3 Energy use at rest

Organ	Resting power (W) of 68 kg individual
Liver	26
Brain	19
Heart	7
Kidneys	11
Skeletal muscle	18
Remainder of body	19
Total	**100**

Energy Use in Activities

Your body stores very little energy as ATP. As your body uses energy, your cells must continuously metabolize carbohydrates, which requires oxygen, as we saw in Equation 11.3. Physiologists can precisely measure the body's energy use by measuring how much oxygen the body is taking up with a respiratory apparatus, as

FIGURE 11.5 The odd masks worn by the fighters measure respiration—the oxygen used by their bodies.



seen in FIGURE 11.5. The device determines the body's *total metabolic energy use*—all of the energy used by the body while performing an activity. This total will include all of the body's basic processes such as breathing plus whatever additional energy is needed to perform the activity. This corresponds to measuring "what you had to pay."

The metabolic energy used in an activity depends on an individual's size, level of fitness, and other variables. But we can make reasonable estimates for the power used in various activities for a typical individual. Some values are given in Table 11.4.

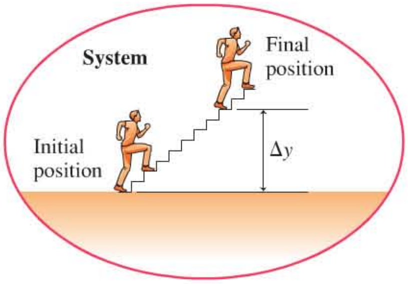
TABLE 11.4 Metabolic power use during activities

Activity	Metabolic power (W) of 68 kg individual
Typing	125
Ballroom dancing	250
Walking at 5 km/h	380
Cycling at 15 km/h	480
Swimming at a fast crawl	800
Running at 15 km/h	1150

Efficiency of the Human Body

Suppose you climb a set of stairs at a constant speed, as in FIGURE 11.6. What is your body's efficiency for this process? In this case, we can do a conservation of energy calculation similar to those in Chapter 10; when you climb stairs, "what you get" is a change in potential energy.

FIGURE 11.6 Climbing a set of stairs.



For such problems, we can adapt the work-energy equation to include only those energy changes that we need to consider. If you climb at a constant speed, there's no change in kinetic energy. And given how we define the system, there is no work—there is no external input of energy, as there would be if you took the elevator. As you climb the stairs, your body uses chemical energy to run your muscles for the climb; chemical energy decreases and potential energy increases. As with other cases we've seen, thermal energy increases as well. This is something you know well: If you climb several sets of stairs, you certainly warm up in the process!

For this case of climbing stairs, the work-energy equation reduces to

$$\Delta E_{\text{chem}} + \Delta U_{\text{g}} + \Delta E_{\text{th}} = 0 \qquad (11.4)$$

Thermal energy and gravitational potential energy are increasing, so ΔE_{th} and ΔU_{g} are positive; chemical energy is being used, so ΔE_{chem} is a negative number. We can get a better feeling about what is happening by rewriting Equation 11.4 as

$$|\Delta E_{\text{chem}}| = \Delta U_{\text{g}} + \Delta E_{\text{th}}$$

The *magnitude* of the change in the chemical energy is equal to the sum of the changes in the gravitational potential and thermal energies. Chemical energy from your body is converted into potential energy and thermal energy; in the final position, you are at a greater height and your body is slightly warmer.

Earlier in the chapter, we noted that the efficiency for stair climbing is about 25%. Let's see where that number comes from.

1. *What you get.* What you get is the change in potential energy: You have raised your body to the top of the stairs. If you climb a flight of stairs of vertical height Δy, we can easily compute the increase in potential energy $\Delta U_{\text{g}} = mg\Delta y$. Assuming a mass of 68 kg and a change in height of 2.7 m (about 9 ft, a reasonable value for a flight of stairs), we compute (to two significant figures)

$$\Delta U_{\text{g}} = (68\text{ kg})(9.8\text{ m/s}^2)(2.7\text{ m}) = 1800\text{ J}$$

2. *What you had to pay.* The cost is the metabolic energy your body used in completing the task. As we've seen, physiologists can measure directly how much energy $|\Delta E_{\text{chem}}|$ your body uses to perform a task. A typical value for climbing a flight of stairs is

$$|E_{\text{chem}}| = 7200\text{ J}$$

Given the definition of efficiency in Equation 11.2, we can compute an efficiency for climbing the stairs:

$$e = \frac{\Delta U_{\text{g}}}{|\Delta E_{\text{chem}}|} = \frac{1800\text{ J}}{7200\text{ J}} = 0.25 = 25\%$$

▶ High heating costs? BIO The daily energy use of mammals is much higher than that of reptiles, largely because mammals use energy to maintain a constant body temperature. A 40 kg timber wolf uses approximately 19,000 kJ during the course of a day. A Komodo dragon, a reptilian predator of the same size, uses only 2100 kJ.

For the types of activities we will consider in this chapter, such as running, walking, and cycling, the body's efficiency is typically in the range of 20–30%. **We will generally use a value of 25% for the body's efficiency for our calculations.** Efficiency varies from individual to individual and from activity to activity, but this rough approximation will be sufficient for our purposes in this chapter.

The metabolic power values given in Table 11.4 represent the energy *used by the body* while these activities are being performed. Given that we assume an efficiency of 25%, the body's actual *useful power output* is quite a bit less than this. The table's value for cycling at 15 km/h (a bit less than 10 mph) is 480 W. If we assume that the efficiency for cycling is 25%, the actual power going to forward propulsion will only be 120 W. An elite racing cyclist whizzing along at 35 km/h is using about 300 W for forward propulsion. This is a surprisingly low figure, as noted in FIGURE 11.7.

The energy you use per second while running is proportional to your speed; running twice as fast takes approximately twice as much power. But running twice as fast takes you twice as far in the same time, so the energy you use to run a certain distance doesn't depend on how fast you run! Running a marathon takes approximately the same amount of energy whether you complete it in 2 hours, 3 hours, or 4 hours; it is only the power that varies.

FIGURE 11.7 Amazingly, a racing cyclist moving at 35 km/h uses about the same power as an electric scooter moving at 5 km/h.

NOTE ▶ It is important to remember the distinction between the metabolic energy used to perform a task and the work done in a physics sense; these values can be quite different. Your muscles use power when applying a force, even when there is no motion. Holding a weight above your head involves no external work, but it clearly takes metabolic power to keep the weight in place! ◀

This distinction—between the work done (what you get) and the energy used by the body (what you had to pay)—is important to keep in mind when you do calculations on energy used by the body. We'll consider two different cases:

- For some tasks, such as climbing stairs, we compute the energy change that is the outcome of the task; that is, we compute what you get. If we assume an efficiency of 25%, the energy used by the body (what you had to pay) is 4 times this amount.
- For other tasks, such as cycling, we use data from metabolic studies (such as data in Table 11.4). This is the actual power used by the body to complete a task—in other words, what you had to pay. If we assume an efficiency of 25%, the useful power output (what you get) is 1/4 of this value.

CONCEPTUAL EXAMPLE 11.3 Energy in weightlifting

A weightlifter lifts a 50 kg bar from the floor to a position over his head and back to the floor again 10 times in succession. At the end of this exercise, what energy transformations have taken place?

REASON We will take the system to be the weightlifter plus the bar. The environment does no work on the system, and we will assume that the time is short enough that no heat is transferred from the system to the environment. The bar has returned to its starting position and is not moving, so there has been no change in potential or kinetic energy. The equation for energy conservation is thus

$$\Delta E_{\text{chem}} + \Delta E_{\text{th}} = 0$$

This equation tells us that ΔE_{chem} must be negative. This makes sense because the muscles *use* chemical energy—depleting the body's store—each time the bar is raised or lowered. Ultimately, all of this energy is transformed into thermal energy.

ASSESS Most exercises in the gym—lifting weights, running on a treadmill—involve only the transformation of chemical energy into thermal energy.

EXAMPLE 11.4 Energy usage for a cyclist

A cyclist pedals for 20 min at a speed of 15 km/h. How much metabolic energy is required? How much energy is used for forward propulsion?

PREPARE Table 11.4 gives a value of 480 W for the power used in cycling at a speed of 15 km/h. 480 W is the power used by the body; the power going into forward propulsion is much less than this. The cyclist uses energy at this rate for 20 min, or 1200 s.

SOLVE We know the power and the time, so we can compute the energy needed by the body as follows:

$$\Delta E = P\Delta t = (480\text{ J/s})(1200\text{ s}) = 580\text{ kJ}$$

If we assume an efficiency of 25%, only 25%, or 140 kJ, of this energy is used for forward propulsion. The remainder goes into thermal energy.

ASSESS How much energy is 580 kJ? A look at Table 11.2 shows that this is slightly more than the amount of energy available in a large apple, and only 10% of the energy available in a large fast-food meal. If you eat such a meal and plan to "work it off" by cycling, you should plan on cycling at a pretty good clip for a bit over 3 h.

EXAMPLE 11.5 How many flights?

How many flights of stairs could you climb on the energy contained in a 12 oz can of soda? Assume that your mass is 68 kg and that a flight of stairs has a vertical height of 2.7 m (9 ft).

PREPARE In Example 11.2 we found that the soda contains 680 kJ of chemical energy. We'll assume that all of this added energy is available to be transformed into the mechanical energy of climbing stairs, at the typical 25% efficiency. We'll also assume you ascend the stairs at constant speed, so your kinetic energy doesn't change. What you get in this case is increased gravitational potential energy, and what you had to pay is the 680 kJ obtained by "burning" the chemical energy of the soda.

SOLVE At 25% efficiency, the amount of chemical energy transformed into increased potential energy is

$$\Delta U_g = (0.25)(680 \times 10^3\text{ J}) = 1.7 \times 10^5\text{ J}$$

Because $\Delta U_g = mg\Delta y$, the height gained is

$$\Delta y = \frac{\Delta U_g}{mg} = \frac{1.7 \times 10^5\text{ J}}{(68\text{ kg})(9.8\text{ m/s}^2)} = 255\text{ m}$$

With each flight of stairs having a height of 2.7 m, the number of flights climbed is

$$\frac{255\text{ m}}{2.7\text{ m}} \cong 94\text{ flights}$$

ASSESS This is almost enough to get to the top of the Empire State Building—all fueled by one can of soda! This is a remarkable result.

Energy Storage

The body gets energy from food; if this energy is not used, it will be stored. A small amount of energy needed for immediate use is stored as ATP. A larger amount of energy is stored as chemical energy of glycogen and glucose in muscle tissue and the liver. A healthy adult might store 400 g of these carbohydrates, which is a little more carbohydrate than is typically consumed in one day.

If the energy input from food continuously exceeds the energy outputs of the body, this energy will be stored in the form of fat under the skin and around the organs. From an energy point of view, gaining weight is simply explained!

EXAMPLE 11.6 Running out of fuel

The body stores about 400 g of carbohydrates. Approximately how far could a 68 kg runner travel on this stored energy?

PREPARE Table 11.1 gives a value of 17 kJ per g of carbohydrate. The 400 g of carbohydrates in the body contain an energy of

$$E_{\text{chem}} = (400\text{ g})(17 \times 10^3\text{ J/g}) = 6.8 \times 10^6\text{ J}$$

SOLVE Table 11.4 gives the power used in running at 15 km/hr as 1150 W. The time that the stored chemical energy will last at this rate is

$$\Delta t = \frac{\Delta E_{\text{chem}}}{P} = \frac{6.8 \times 10^6\text{ J}}{1150\text{ W}} = 5.91 \times 10^3\text{ s} = 1.64\text{ h}$$

And the distance that can be covered during this time at 15 km/h is

$$\Delta x = v\Delta t = (15\text{ km/h})(1.64\text{ h}) = 25\text{ km}$$

to two significant figures.

ASSESS A marathon is longer than this—just over 42 km. Even with "carbo loading" before the event (eating high-carbohydrate meals), many marathon runners "hit the wall" before the end of the race as they reach the point where they have exhausted their store of carbohydrates. The body has other energy stores (in fats, for instance), but the rate that they can be drawn on is much lower.

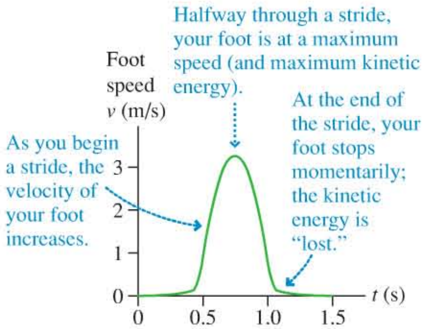
Energy and Locomotion

When you walk at a constant speed on level ground, your kinetic energy is constant. Your potential energy is also constant. So why does your body need energy to walk? Where does this energy go?

We use energy to walk because of mechanical inefficiencies in our gait. FIGURE 11.8 shows how the speed of your foot typically changes during each stride. The kinetic energy of your leg and foot increases, only to go to zero at the end of the stride. The kinetic energy is mostly transformed into thermal energy in your muscles and in your shoe. This thermal energy is lost; it can't be used for making more strides.

Footwear can be designed to minimize the loss of kinetic energy to thermal energy. A spring in the sole of the shoe can store potential energy, which can be returned to kinetic energy during the next stride. Such a spring will make the collision with the ground more elastic. We saw in Chapter 10 that the tendons in the ankle do store a certain amount of energy during a stride; very stout tendons in the legs of kangaroos store energy even more efficiently. Their peculiar hopping gait is efficient at high speeds.

FIGURE 11.8 Human locomotion analysis.



Where do you wear the weights? BIO If you wear a backpack with a mass equal to 1% of your body mass, your energy expenditure for walking will increase by 1%. But if you wear ankle weights with a combined mass of 1% of your body mass, the increase in energy expenditure is 6%, because you must repeatedly accelerate this extra mass. If you want to "burn more fat," wear the weights on your ankles, not on your back! If you are a runner who wants to shave seconds off your time in the mile, you might try lighter shoes.

STOP TO THINK 11.2 A runner is moving at a constant speed on level ground. Chemical energy in the runner's body is being transformed into other forms of energy; most of the chemical energy is transformed into

A. Kinetic energy. B. Potential energy. C. Thermal energy.

11.4 Thermal Energy and Temperature

We have frequently spoken of the energy that is transformed into thermal energy as being "lost." Regardless of whether "the system" is a car, a power plant, or your body, this thermal energy is simply exhausted into the environment. But why isn't thermal energy in your body and other systems converted to other forms of energy and used for practical purposes? To continue our study of how energy is used, we need to understand a bit more about one particular kind of energy—thermal energy.

What do you mean when you say something is "hot"? Do you mean that it has a high temperature? Or do you mean that it has a lot of thermal energy? Are both of these definitions the same? Let's give some thought to the definitions of temperature and thermal energy, and the relationship between them.

An Atomic View of Thermal Energy and Temperature

Consider the simplest possible atomic system, the **ideal gas** of atoms seen in FIGURE 11.9. In Chapter 10, we defined thermal energy to be the energy associated with the motion of the atoms and molecules that make up an object. Because there are no molecular bonds, the atoms in an ideal gas have only the kinetic energy of their motion. Thus, **the thermal energy of an ideal gas is equal to the total kinetic energy of the moving atoms in the gas.**

If you take a container of an ideal gas and place it over a flame, as in FIGURE 11.10 on the next page, energy will be transferred from the hot flame to the cooler gas. This is a new type of energy transfer, one we'll call **heat**. We will define the nature of heat later in the chapter; for now, we'll focus on the changes in the gas when it is heated.

FIGURE 11.9 Motion of atoms in an ideal gas.

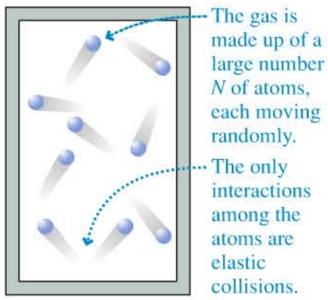
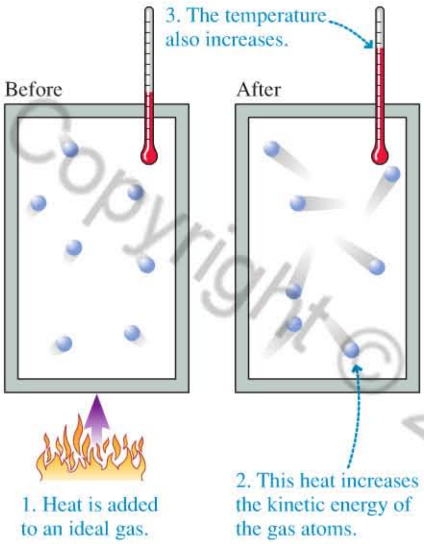


FIGURE 11.10 Heating an ideal gas.



As Figure 11.10 shows, heating the gas causes the atoms to move faster and thus increases the thermal energy of the gas. Heating the gas also increases its temperature. The observation that both increase suggests that temperature must be related to the kinetic energy of the gas atoms, but how? We can get an important hint from the following fact: The temperature of a system does not depend on the size of the system. If you mix together two glasses of water, each at a temperature of 20°C, you will have a larger volume of water at the same temperature of 20°C. The combined volume has more atoms, and therefore more *total* thermal energy, but each atom is moving about just as it was before, and so the *average* kinetic energy per atom is unchanged. It is this average kinetic energy of the atoms that is related to the temperature. **The temperature of an ideal gas is a measure of the *average* kinetic energy of the atoms that make up the gas.**

We'll formalize and extend these ideas in Chapter 12, but, for now, we'll be able to say quite a bit about temperature and thermal energy by referring to the properties of this simple model.

Temperature Scales

Temperature, in an ideal gas, is related to the average kinetic energy of atoms. We can make a similar statement about other materials, be they solid, liquid, or gas: The higher the temperature, the faster the atoms move. But how do you measure temperature? The common glass-tube thermometer works by allowing a small volume of mercury or alcohol to expand or contract when placed in contact with a "hot" or "cold" object. Other thermometers work in different ways, but all—at a microscopic level—depend on the speed of the object's atoms as they collide with the atoms in the thermometer. That is, all thermometers are sampling the average kinetic energy of the atoms in the object.

Thermal expansion of the liquid in the thermometer pushes it higher when immersed in hot water than in ice water.

A thermometer needs a temperature scale to be a useful measuring device. The scale used in scientific work (and in almost every country in the world) is the *Celsius scale*. As you likely know, the Celsius scale is defined so that the freezing point of water is 0°C and the boiling point is 100°C. The units of the Celsius temperature scale are "degrees Celsius," which we abbreviate as °C. The Fahrenheit scale, still widely used in the United States, is related to the Celsius scale by

$$T(°\text{C}) = \frac{5}{9}\left(T(°\text{F}) - 32°\right) \qquad T(°\text{F}) = \frac{9}{5}T(°\text{C}) + 32° \tag{11.5}$$

Both the Celsius and the Fahrenheit scales have a zero point that is arbitrary—simply an agreed-upon convention—and both allow negative temperatures. If, instead, we use the average kinetic energy of ideal-gas atoms as our basis for the definition of temperature, our temperature scale will have a natural zero—the point at which kinetic energy is zero. Kinetic energy is always positive, so the zero on our temperature scale will be an **absolute zero;** no temperature below this is possible.

This is how zero is defined on the temperature scale called the *Kelvin scale:* **Zero degrees is the point at which the kinetic energy of atoms is zero.** All temperatures on the Kelvin scale are positive, so it is often called an *absolute temperature scale.* The units of the Kelvin temperature scale are "kelvin" (not degrees kelvin!), abbreviated K.

The spacing between divisions on the Kelvin scale is the same as that of the Celsius scale; the only difference is the position of the zero point. Absolute zero—the temperature at which atoms would cease moving—is -273°C. The conversion between Celsius and Kelvin temperatures is therefore quite straightforward:

$$T(\text{K}) = T(°\text{C}) + 273 \qquad T(°\text{C}) = T(\text{K}) - 273 \tag{11.6}$$

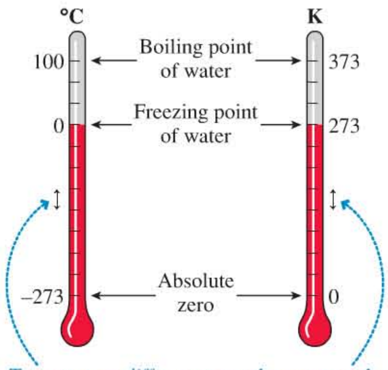
The size of a degree on the Kelvin and Celsius scales is the same. This means that a temperature *difference* is the same on both scales:

$$\Delta T(\text{K}) = \Delta T(°\text{C})$$

On the Kelvin scale, the freezing point of water at 0°C is $T = 0 + 273 = 273$ K. A 30°C warm summer day is $T = 303$ K on the Kelvin scale. **FIGURE 11.11** gives a side-by-side comparison of these scales.

NOTE ► From now on, we will use the symbol T for temperature in kelvin. We will denote other scales by showing the units in parentheses. In the equations in this chapter and the rest of the text, **T must be interpreted as a temperature in kelvin.** ◄

FIGURE 11.11 Celsius and Kelvin temperature scales.



Temperature *differences* are the same on the Celsius and Kelvin scales. The temperature difference between the freezing point and boiling point of water is 100°C or 100 K.

EXAMPLE 11.7 **Temperature scales**

The coldest temperature ever measured on earth was −129°F, in Antarctica. What is this in °C and K?

SOLVE We use Equation 11.5 to convert the temperature to the Celsius scale:

$$T(°\mathrm{C}) = \frac{5}{9}(-129° - 32°) = -89°\mathrm{C}$$

We can then use Equation 11.6 to convert this to kelvin:

$$T = -89 + 273 = 184 \text{ K}$$

ASSESS This is cold, but quite a bit warmer than the coldest temperatures achieved in the laboratory.

Relating Temperature and Thermal Energy

We noted above that the temperature of an ideal gas is a measure of the average kinetic energy of the atoms. It can be shown that temperature on the Kelvin scale is related to the average kinetic energy per atom by

$$T = \frac{2}{3}\frac{K_{\text{avg}}}{k_{\text{B}}} \tag{11.7}$$

where k_{B} is a constant known as **Boltzmann's constant.** Its value is

$$k_{\text{B}} = 1.38 \times 10^{-23} \text{ J/K}$$

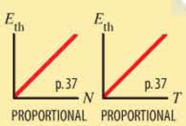
We can rearrange Equation 11.7 to give the average kinetic energy in terms of the temperature:

$$K_{\text{avg}} = \frac{3}{2}k_{\text{B}}T \tag{11.8}$$

The thermal energy of an ideal gas consisting of N atoms is the sum of the kinetic energies of the individual atoms:

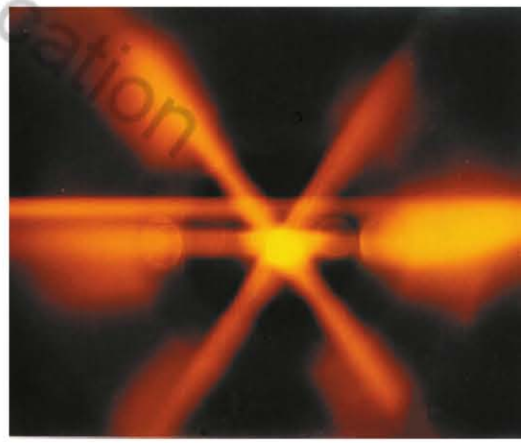
$$E_{\text{th}} = NK_{\text{avg}} = \frac{3}{2}Nk_{\text{B}}T \tag{11.9}$$

Thermal energy of an ideal gas of N atoms



For an ideal gas, **thermal energy is directly proportional to temperature.** Consequently, a change in the thermal energy of an ideal gas is proportional to a change in temperature:

$$\Delta E_{\text{th}} = \frac{3}{2}Nk_{\text{B}}\Delta T \tag{11.10}$$



Optical molasses It isn't possible to reach absolute zero, where the atoms would be still, but it is possible to get quite close by slowing the atoms down directly. These crossed laser beams produce what is known as "optical molasses." As we will see in Chapter 28, light is made of photons, which carry energy and momentum. Interactions of the atoms of a diffuse gas with the photons cause the atoms to slow down. In this manner atoms can be slowed to speeds that correspond to a temperature as cold as 5×10^{-10} K!

ActivPhysics Online 8.1, 8.2, 8.3

Is it cold in space? The Space Shuttle orbits in the upper thermosphere, about 300 km above the surface of the earth. There is still a trace of atmosphere left at this altitude, and it has quite a high temperature—over 1000°C. Although the average speed of the air molecules here is high, there are so few air molecules present that the thermal energy is extremely low.

This relationship between a change in temperature and a change in thermal energy is for an ideal gas, but solids, liquids, and other gases all follow similar rules, as we will see in the next chapter. For now, we will simply note two important conclusions that apply to any substance:

1. The thermal energy of a substance is proportional to the number of atoms. A gas with more atoms has more thermal energy than a gas at the same temperature with fewer atoms.
2. A change in temperature causes a proportional change in the substance's thermal energy. A larger temperature change causes a larger change in thermal energy.

EXAMPLE 11.8 Energy needed to warm up a room

A large bedroom contains about 1×10^{27} molecules of air. (In the next chapter, we'll see how to calculate this.) Estimate the energy required to raise the temperature of the air in the room by 5°C.

PREPARE We'll model the air as an ideal gas. Equation 11.10 relates the change in thermal energy of an ideal gas to a change in temperature. The actual temperature of the gas doesn't matter—only the change. The temperature increase is given as 5°C, implying a change in the absolute temperature by the same amount: $\Delta T = 5\text{ K}$.

SOLVE We can use Equation 11.10 to calculate the amount by which the room's thermal energy must be increased:

$$\Delta E_{\text{th}} = \frac{3}{2}Nk_B\Delta T = \frac{3}{2}(1 \times 10^{27})(1.38 \times 10^{-23}\text{ J/K})(5\text{ K}) = 1 \times 10^5\text{ J} = 100\text{ kJ}$$

This is the energy we would have to supply—probably in the form of heat from a furnace—to raise the temperature.

ASSESS 100 kJ isn't that much energy. Table 11.2 shows it to be less than the food energy in a carrot! This seems reasonable because you know that your furnace can quickly warm up the air in a room. Heating up the walls and furnishings is another story.

STOP TO THINK 11.3 Two samples of ideal gas, sample 1 and sample 2, have the same thermal energy. Sample 1 has twice as many atoms as sample 2. What can we say about the temperatures of the two samples?

A. $T_1 > T_2$ B. $T_1 = T_2$ C. $T_1 < T_2$

11.5 Heat and the First Law of Thermodynamics

In Chapter 10, we saw that a system could exchange energy with the environment through two different means: work and heat. Work was treated in some detail in Chapter 10; now it is time to look at the transfer of energy by heat. This will begin our study of a topic called *thermodynamics,* the study of thermal energy and heat and their relationships to other forms of energy and energy transfer.

What Is Heat?

Heat is a more elusive concept than work. We use the word "heat" very loosely in the English language, often as synonymous with "hot." We might say, on a very hot day, "This heat is oppressive." If your apartment is cold, you may say, "Turn up the heat." It's time to develop more precise language to discuss these concepts.

Suppose you put a pan of cold water on the stove. If you light the burner so that there is a hot flame under the pan, as in FIGURE 11.12a, the temperature of the water increases. We've added energy to the system—the pan of water—as heat. You know, from everyday experience, that heat always flows "downhill" in the sense that energy is transferred from a hotter object to a colder object. If there is no temperature difference, no energy is transferred. We'll explore this idea further, but, for now, we can use this familiar observation as a definition of heat: **Heat is energy transferred between two objects at different temperatures.**

FIGURE 11.12 Two different means of raising the temperature of a pan of water.

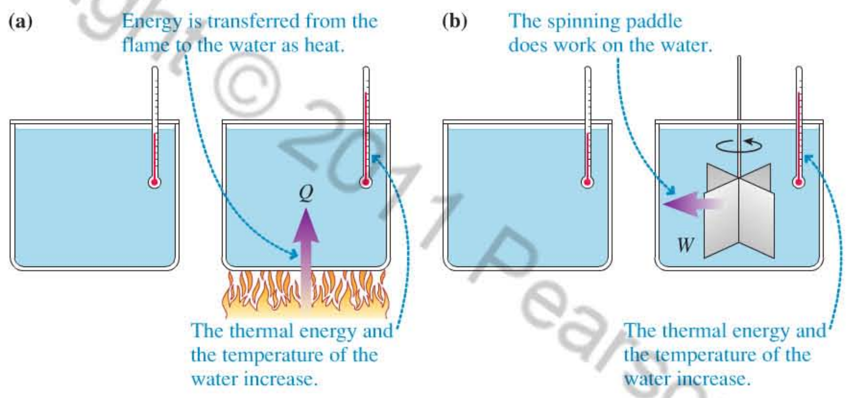


FIGURE 11.12b shows the same system, but this time work is being done on the system by means of a rapidly spinning paddle. Careful experiments by the British physicist James Joule in the 1840s found that doing work on the system also increases the temperature. In fact, the temperature increase and the final state of the system are exactly the same regardless of whether energy is added as heat or an equal amount of energy is added by doing work. This implies that heat and work are in some sense equivalent: **Heat and work are simply two different ways of transferring energy to or from a system.**

An Atomic Model of Heat

Let's consider an atomic model to explain why thermal energy is transferred from higher temperatures to lower. FIGURE 11.13 shows a rigid, insulated container that is divided into two sections by a very thin membrane. Each side is filled with a gas of the same kind of atoms. The left side, which we'll call system 1, is at an initial temperature T_{1i}. System 2 on the right is at an initial temperature T_{2i}. We imagine the membrane to be so thin that atoms can collide at the boundary as if the membrane were not there, yet it is a barrier that prevents atoms from moving from one side to the other.

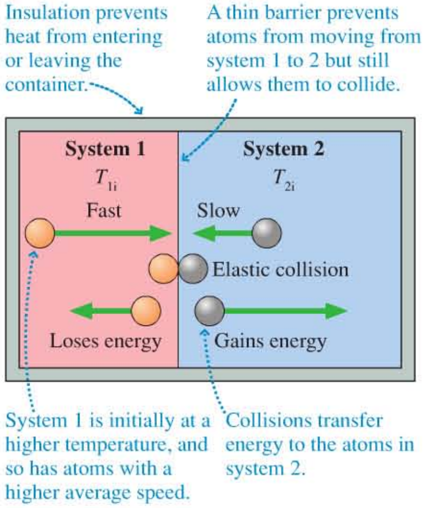
Suppose that system 1 is initially at a higher temperature: $T_{1i} > T_{2i}$. This means that the atoms in system 1 have a higher average kinetic energy. Figure 11.13 shows a fast atom and a slow atom approaching the barrier from opposite sides. They undergo a perfectly elastic collision at the barrier. Although no net energy is lost in a perfectly elastic collision, the faster atom loses energy while the slower one gains energy. In other words, there is an energy *transfer* from the faster atom's side to the slower atom's side.

Because the atoms in system 1 are, on average, more energetic than the atoms in system 2, *on average* the collisions transfer energy from system 1 to system 2. This is not true for every collision. Sometimes a fast atom in system 2 collides with a slow atom in system 1, transferring energy from 2 to 1. But the net energy transfer, from all collisions, is from the warmer system 1 to the cooler system 2. This transfer of energy is heat; **thermal energy is transferred from the faster moving atoms on the warmer side to the slower moving atoms on the cooler side.**

TRY IT YOURSELF

Energetic cooking You can do a modern version of Joule's experiment in the kitchen. Next time you mix food in a blender, notice that it actually warms the food. As the blades rotate, friction increases the thermal energy of the food in the blender. A blender uses about as much power as a microwave, and most of this energy ends up as thermal energy in the food, so the temperature rises as the food is chopped and blended.

FIGURE 11.13 Collisions at a barrier transfer energy from faster molecules to slower molecules.



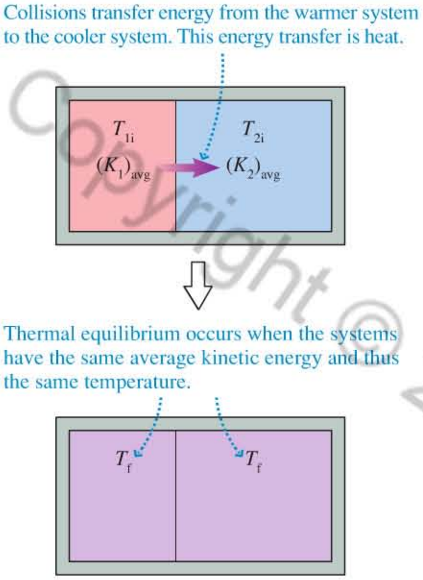


FIGURE 11.14 Two systems in thermal contact exchange thermal energy.

This transfer will continue until a stable situation is reached. This is a situation we call **thermal equilibrium.** How do the systems "know" when they've reached thermal equilibrium? Energy transfer continues until the atoms on both sides of the barrier have the *same average kinetic energy.* Once the average kinetic energies are the same, individual collisions will still transfer energy from one side to the other. But since both sides have atoms with the same average kinetic energies, the amount of energy transferred from 1 to 2 will equal that transferred from 2 to 1. Once the average kinetic energies are the same, there will be no more net energy transfer.

As we've seen, the average kinetic energy of the atoms in a system is directly proportional to the system's temperature. If two systems exchange energy until their atoms have the same average kinetic energy, we can say that

$$T_{1f} = T_{2f} = T_f$$

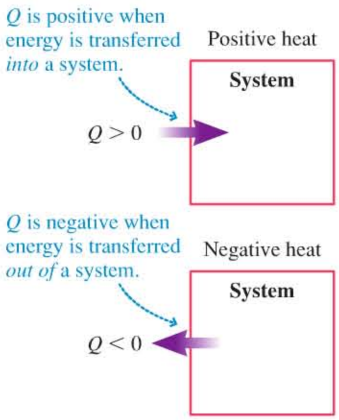
That is, heat is transferred until the two systems reach a common final temperature; we call this final state **thermal equilibrium.** We considered a rather artificial system in this case, but the result is quite general: **Two systems placed in thermal contact will transfer thermal energy from hot to cold until their final temperatures are the same.** This process is illustrated in FIGURE 11.14.

Heat is a transfer of energy. The sign that we use for transfers is defined in FIGURE 11.15. In the process of Figure 11.14, Q_1 is negative because system 1 loses energy; Q_2 is positive because system 2 gains energy. No energy escapes from the container, so all of the energy that was lost by system 1 was gained by system 2. We can write that as:

$$Q_2 = -Q_1$$

The heat energy lost by one system is gained by the other.

FIGURE 11.15 The sign of Q.



The First Law of Thermodynamics

We need to broaden our work-energy equation, Equation 11.1, to include energy transfers in the form of heat. The sign conventions for heat, as shown in Figure 11.15, have the same form as those for work—a positive value means a transfer into the system, a negative value means a transfer out of the system. Thus, when we include heat Q in the work-energy equation, it appears on the right side of the equation along with work W:

$$\Delta K + \Delta U + \Delta E_{th} + \Delta E_{chem} + \cdots = W + Q \qquad (11.11)$$

This equation is our most general statement to date about energy and the conservation of energy; it includes all of the energy transfers and transformations we have discussed.

In Chapter 10, we focused on systems where the potential and kinetic energies could change, such as a sled moving down a hill. Earlier in this chapter, we looked at the body, where the chemical energy changes. Now let's consider systems in which only the thermal energy changes. That is, we will consider systems that aren't moving, that aren't changing chemically, but whose temperatures can change. Such systems are the province of what is called **thermodynamics.** The question of how to keep your house cool in the summer is a question of thermodynamics. If energy is transferred into your house, the thermal energy increases and the temperature rises. To reduce the temperature, you must transfer energy out of the house. This is the purpose of an air conditioner, as we'll see in a later section.

If we consider cases in which only thermal energy changes, Equation 11.11 can be simplified. This simpler version is a statement of conservation of energy for systems in which only the thermal energy changes; we call it the **first law of thermodynamics:**

First law of thermodynamics For systems in which only the thermal energy changes, the change in thermal energy is equal to the energy transferred into or out of the system as work W, heat Q, or both:

$$\Delta E_{th} = W + Q \tag{11.12}$$

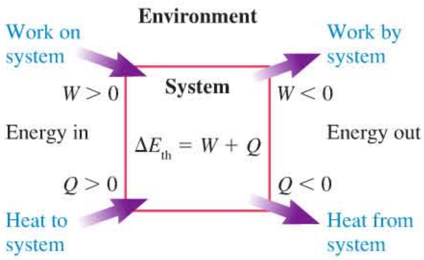


FIGURE 11.16 Energy transfers in a thermodynamic system.

Only work and heat, two ways of transferring energy between a system and the environment, cause the system's energy to change. In thermodynamic systems, the only energy change will be a change in thermal energy. Whether this energy increases or decreases depends on the signs of W and Q, as we've seen. The possible energy transfers between a system and the environment are illustrated in **FIGURE 11.16**.

CONCEPTUAL EXAMPLE 11.9 **Compressing a gas**

Suppose a gas is in an insulated container, so that no heat energy can escape. If a piston is used to compress the gas, what happens to the temperature of the gas?

REASON The piston applies a force to the gas, and there is a displacement. This means that work is done on the gas by the piston ($W > 0$). No thermal energy can be exchanged with the environment, meaning $Q = 0$. Since energy is transferred into the system, the thermal energy of the gas must increase. This means that the temperature must increase as well.

ASSESS This result makes sense in terms of everyday observations you may have made. When you use a bike pump to inflate a tire, the pump and the tire get warm. This temperature increase is largely due to the warming of the air by the compression.

EXAMPLE 11.10 **Work and heat in an ideal gas**

Suppose an uninsulated container holds 5.0×10^{22} molecules of an ideal gas. 50 J of work are done on the gas by a piston that compresses it. The temperature of the gas increases by 30°C during this process. How much heat is transferred to or from the environment?

PREPARE This is an energy conservation problem. We can first use the temperature increase to determine the change in thermal energy of the gas; then use this to determine how much heat energy goes in or out.

SOLVE For an ideal gas, the change in thermal energy is given by Equation 11.10:

$$\Delta E_{th} = \frac{3}{2} N k_B \Delta T$$
$$= \left(\frac{3}{2}\right)(5.0 \times 10^{22})(1.38 \times 10^{-23}\ \text{J/K})(30\ \text{K}) = 31\ \text{J}$$

The first law of thermodynamics, Equation 11.12, tells us that the change in thermal energy ΔE_{th} of the gas is the sum of W and Q, the total energy added to the system. W is known, and we have just calculated ΔE_{th}. Combining, we get

$$Q = \Delta E_{th} - W = 31\ \text{J} - 50\ \text{J} = -19\ \text{J}$$

ASSESS Q is negative, meaning that energy is transferred outward from the hot gas to the cooler environment in this process.

Energy-Transfer Diagrams

Suppose you drop a hot rock into the ocean. Heat is transferred from the rock to the ocean until the rock and ocean are the same temperature. Although the ocean warms up ever so slightly, ΔT_{ocean} is so small as to be completely insignificant.

An **energy reservoir** is an object or a part of the environment so large that, like the ocean, its temperature does not noticeably change when heat is transferred between the system and the reservoir. A reservoir at a higher temperature than the system is called a *hot reservoir*. A vigorously burning flame is a hot reservoir for small objects placed in the flame. A reservoir at a lower temperature than the system is called a *cold reservoir*. The ocean is a cold reservoir for the hot rock. We will use T_H and T_C to designate the temperatures of the hot and cold reservoirs.

Heat energy is transferred between a system and a reservoir if they have different temperatures. We will define

$$Q_H = \text{amount of heat transferred to or from a hot reservoir}$$

$$Q_C = \text{amount of heat transferred to or from a cold reservoir}$$

By definition, Q_H and Q_C are *positive* quantities.

FIGURE 11.17a shows a heavy copper bar placed between a hot reservoir (at temperature T_H) and a cold reservoir (at temperature T_C). Heat Q_H is transferred from the hot

FIGURE 11.17 Energy-transfer diagrams.

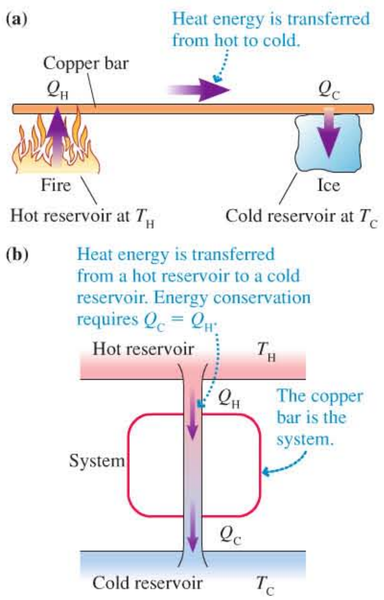
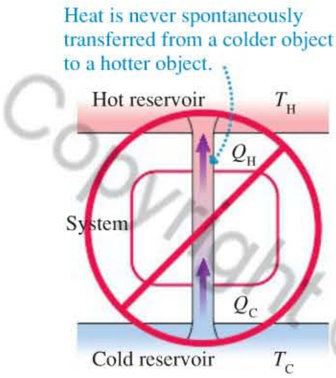


FIGURE 11.18 An impossible energy transfer.



reservoir into the copper, and heat Q_C is transferred from the copper to the cold reservoir. FIGURE 11.17b is an **energy-transfer diagram** for this process. The hot reservoir is generally drawn at the top, the cold reservoir at the bottom, and the system—the copper bar in this case—between them. The reservoirs and the system are connected by "pipes" that show the energy transfers. Figure 11.17b shows heat Q_H being transferred into the system and Q_C being transferred out.

FIGURE 11.18 illustrates an important fact about heat transfers that we have discussed: Spontaneous transfers go in one direction only, from hot to cold. This is an important result that has significant practical implications.

CONCEPTUAL EXAMPLE 11.11 **Energy transfers and the body**

Why—in physics terms—is it more taxing on the body to exercise in very hot weather?

REASON Your body continuously converts chemical energy to thermal energy, as we have seen. In order to maintain a constant body temperature, your body must continuously transfer heat to the environment. This is a simple matter in cool weather when heat is spontaneously transferred to the environment, but when the air temperature is higher than your body temperature, your body cannot cool itself this way and must use other mechanisms to transfer this energy, such as perspiring. These mechanisms require additional energy expenditure.

ASSESS Strenuous exercise in hot weather can easily lead to a rise in body temperature if the body cannot exhaust heat quickly enough.

Falling water turns a waterwheel.

STOP TO THINK 11.4 You have driven your car for a while and now turn off the engine. Your car's radiator is at a higher temperature than the air around it. Considering the radiator as the system, we can say that

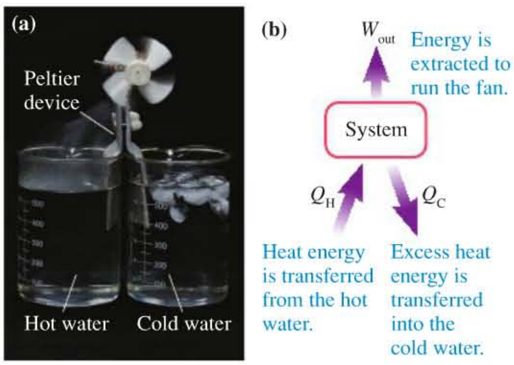
A. $Q > 0$ B. $Q = 0$ C. $Q < 0$

11.6 Heat Engines

In the early stages of the industrial revolution, most of the energy needed to run mills and factories came from water power. Water in a high reservoir will naturally flow downhill. A waterwheel can be used to harness this natural flow of water to produce some useful energy because some of the potential energy lost by the water as it flows downhill can be converted into other forms.

It is possible to do something similar with heat. Thermal energy is naturally transferred from a hot reservoir to a cold reservoir; it is possible to take some of this energy as it is transferred and convert it to other forms. This is the job of a device known as a **heat engine.**

FIGURE 11.19 A simple heat engine.



A simple example of a heat engine is shown in FIGURE 11.19a. Here a *Peltier device*—a device that produces a voltage if there is a temperature difference between its two sides—is sandwiched between two aluminum vanes that sit in cups of water. The cup of water on the left is hot; the cup on the right is cold. The temperature difference produces a voltage that drives the fan.

FIGURE 11.19b is a diagram that illustrates the operation of this device in thermodynamic terms. The Peltier device is the system. Heat energy is transferred through the device from the hot water to the cold water. As the transfer occurs, some of this energy is transformed into electric energy to run the fan.

The energy-transfer diagram of FIGURE 11.20 on the next page illustrates the basic physics of a heat engine. It takes in energy as heat from the hot reservoir, turns some of it into useful work, and exhausts the balance as waste heat in the cold reservoir. Any heat engine has exactly the same schematic.

Efficiency of a Heat Engine

We assume, for a heat engine, that the engine's thermal energy doesn't change. This means that there is no net energy transfer into or out of the heat engine. Because energy is conserved, we can say that the useful work extracted is equal to the difference between the heat energy transferred from the hot reservoir and the heat exhausted into the cold reservoir:

$$W_{\text{out}} = Q_{\text{H}} - Q_{\text{C}}$$

The energy input to the engine is Q_{H} and the energy output is W_{out}.

NOTE ► We earlier defined Q_{H} and Q_{C} to be positive quantities. We also define the energy output of a heat engine W_{out} to be a positive quantity. For heat engines, the directions of the transfers will always be clear, and we will take all of these basic quantities to be positive. ◄

We can use the definition of efficiency, from earlier in the chapter, to compute the heat engine's efficiency:

$$e = \frac{\text{what you get}}{\text{what you had to pay}} = \frac{W_{\text{out}}}{Q_{\text{H}}} = \frac{Q_{\text{H}} - Q_{\text{C}}}{Q_{\text{H}}} \tag{11.13}$$

Q_{H} is what you had to pay because this is the energy of the fuel burned—and, quite literally, paid for—to provide the high temperature of the hot reservoir. The heat energy that is not converted to work ends up in the cold reservoir as waste heat.

Why should we waste energy this way? Why don't we make a heat engine like the one shown in **FIGURE 11.21** that converts 100% of the heat into useful work? The surprising answer is that we can't. **No heat engine can operate without exhausting some fraction of the heat into a cold reservoir**. This isn't a limitation on our engineering abilities. As we'll see, it's a fundamental law of nature.

The maximum possible efficiency of a heat engine is fixed by the *second law of thermodynamics,* which we will explore in detail in Section 11.8. We will not do a detailed derivation, but simply note that the second law gives the maximum efficiency of a heat engine as

$$e_{\text{max}} = 1 - \frac{T_{\text{C}}}{T_{\text{H}}} \tag{11.14}$$

Theoretical maximum efficiency of a heat engine

The maximum efficiency of any heat engine is therefore fixed by the ratio of the temperatures of the hot and cold reservoirs. It is possible to increase the efficiency of a heat engine by increasing the temperature of the hot reservoir or decreasing the temperature of the cold reservoir. The efficiency of Equation 11.14 is also called the *Carnot efficiency,* after a particular heat engine that achieves this maximum possible efficiency. The actual efficiency of real heat engines is usually much less than the theoretical maximum.

NOTE ► The temperatures in Equation 11.14 *must* be in K, not °C. ◄

Because $e_{\text{max}} < 1$, the work done is always less than the heat input ($W_{\text{out}} < Q_{\text{H}}$). Consequently, there *must* be heat Q_{C} exhausted to the cold reservoir. That is why the heat engine of Figure 11.21 is impossible. Think of a heat engine as analogous to a water wheel: The engine "siphons off" some of the energy that is spontaneously flowing from hot to cold, but it can't completely shut off the flow.

The fact that heat engine efficiency is limited—and that heat must be exhausted into a cold reservoir—is of tremendous practical importance. Most of the energy that you use daily comes from the conversion of chemical energy into thermal energy and the subsequent conversion of that energy into other forms. Let's look at some common examples of heat engines.

FIGURE 11.20 The operation of a heat engine.

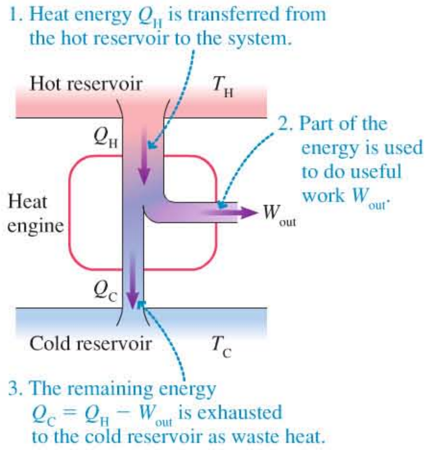
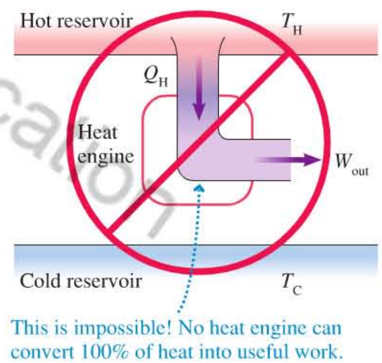
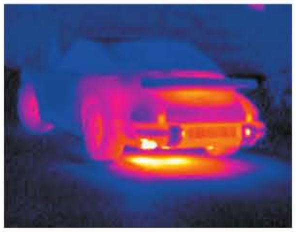


FIGURE 11.21 A perfect (and impossible!) heat engine.



Heat engines

Most of the electricity that you use was generated by heat engines. Coal or other fossil fuels are burned to produce high-temperature, high-pressure steam. The steam does work by spinning a turbine attached to a generator, which produces electricity. Some of the energy of the steam is extracted in this way, but more than half simply flows "downhill" and is deposited in a cold reservoir, often a lake or a river.



Your car gets the energy it needs to run from the chemical energy in gasoline. The gasoline is burned; the resulting hot gases are the hot reservoir. Some of the thermal energy is converted into the kinetic energy of the moving vehicle, but more than 90% is lost as heat to the surrounding air via the radiator and the exhaust, as shown in this thermogram.

There are many small, simple heat engines that are part of things you use daily. This fan, which can be put on top of a wood stove, uses the thermal energy of the stove to provide power to drive air around the room. Where are the hot and cold reservoirs in this device?

EXAMPLE 11.12 **The efficiency of a nuclear power plant**

Energy from nuclear reactions in the core of a nuclear reactor produces high-pressure steam at a temperature of 290°C. After the steam is used to spin a turbine, it is condensed (by using cooling water from a nearby river) back to water at 20°C. The excess heat is deposited in the river. The water is then reheated, and the cycle beings again. What is the maximum possible efficiency that this plant could achieve?

PREPARE A nuclear power plant is a heat engine, with energy transfers as illustrated in Figure 11.20. Q_H is the heat energy transferred to the steam in the reactor core. T_H is the temperature of the steam, 290°C. The steam is cooled and condensed, and the heat Q_C is exhausted to the river. The river is the cold reservoir, so T_C is 20°C.

In kelvin, these temperatures are

$$T_H = 290°\text{C} = 563\text{ K} \qquad T_C = 20°\text{C} = 293\text{ K}$$

SOLVE We use Equation 11.14 to compute the maximum possible efficiency:

$$e_{max} = 1 - \frac{T_C}{T_H} = 1 - \frac{293\text{ K}}{563\text{ K}} = 0.479 \approx 48\%$$

ASSESS This is the maximum possible efficiency. There are practical limitations as well that limit real power plants, whether nuclear or coal- or gas-fired, to an efficiency $e \approx 0.35$. This means that 65% of the energy from the fuel is exhausted as waste heat into a river or lake, where it may cause problematic warming in the local environment.

Heat engines take energy from a hot reservoir, transform some into useful forms of energy, such as mechanical or electric energy, and deposit the rest as thermal energy in a cold reservoir. The laws of physics say that some energy must be deposited in the cold reservoir, but this energy need not be "wasted." Those of you who live in cold climates use the heat energy from your car's engine to warm the car's interior—heat that would otherwise simply be deposited in the environment. There is no cost to warm your car's interior on a chilly day; you are simply putting this "waste" heat to good use. Many college campuses have combined heat and power systems in which the exhausted heat from an electric power plant is used to warm campus buildings. This scheme is quite common in cities in Europe and, increasingly, the United States.

STOP TO THINK 11.5 Which of the following changes (there may be more than one) would increase the maximum theoretical efficiency of a heat engine?

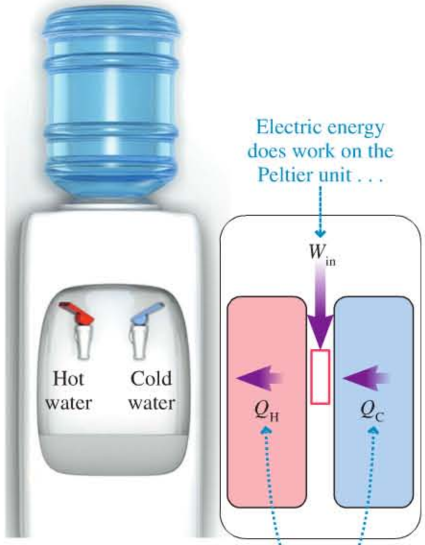
A. Increase T_H B. Increase T_C C. Decrease T_H D. Decrease T_C

11.7 Heat Pumps

The inside of your refrigerator is colder than the air in your kitchen, so heat energy will be transferred from the room to the inside of the refrigerator, warming it. This happens every time you open the door and as heat "leaks" through the walls. But you want the inside of the refrigerator to stay cool. To keep it cool, you need some way to move this heat back out to the warmer room. Transferring heat energy from a cold reservoir to a hot reservoir—the opposite of the natural direction—is the job of a **heat pump.**

The Peltier device that we saw in the last section can also be used as a heat pump. If an electric current is put through the device, a temperature difference develops between the two sides. As time goes by, the cold side gets colder and the hot side gets hotter. A practical application of such a heat pump is the water cooler/heater shown in FIGURE 11.22. The energy that is removed from the cold water ends up as increased thermal energy in the hot water. A single process cools the cold water and heats the hot water.

FIGURE 11.22 A heat pump provides hot and cold water.



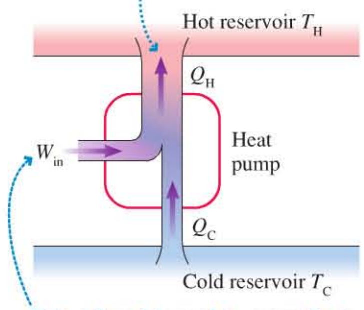
The heat pump in a refrigerator transfers heat from the cold air *inside* the refrigerator to the warmer air in the room. Coils inside the refrigerator that are colder than the inside air take in thermal energy. This energy is transferred to warm coils *outside* the refrigerator that transfer heat to the room. In the antique refrigerator shown on the next page, the cold coils are in the metal bracket at the top of the cabinet, the warm coils are outside the cabinet in the cylindrical unit on top. When the refrigerator is running, these top coils are quite warm. If you look closely at your refrigerator, you will find warm coils or a warm air exhaust that transfers heat to the room, heat that was removed from the inside.

An air conditioner works similarly, transferring heat from the cool air of a house (or a car) to the warmer air outside. You have probably seen room air conditioners that are single units meant to fit in a window. The side of the air conditioner that faces the room is the cool side; the other side of the air conditioner is warm. Heat is pumped from the cool side to the warm side, cooling the room.

You can also use a heat pump to *warm* your house in the winter by moving thermal energy from outside your house to inside. A unit outside the house takes in heat energy that is pumped to a unit inside the house, warming the inside air.

In all of these cases, we are moving energy against the natural direction it would flow. This requires an energy input—work must be done—as shown in the energy–transfer diagram of a heat pump in FIGURE 11.23. Energy must be conserved, so the heat deposited in the hot side must equal the sum of the heat removed from the cold side and the work input:

FIGURE 11.23 The operation of a heat pump.



$$Q_H = Q_C + W_{in}$$

For heat pumps, rather than compute efficiency, we compute an analogous quantity called the **coefficient of performance (COP).** There are two different ways that one can use a heat pump, as noted above. A refrigerator uses a heat pump for cooling, removing heat from a cold reservoir to keep it cold. As Figure 11.23 shows, we must do work to make this happen.

If we use a heat pump for cooling, we define the coefficient of performance as

$$\text{COP} = \frac{\text{what you get}}{\text{what you had to pay}} = \frac{\text{energy removed from the cold reservoir}}{\text{work required to perform the transfer}} = \frac{Q_C}{W_{in}}$$

The second law of thermodynamics limits the efficiency of a heat pump just as it limits the efficiency of a heat engine. The maximum possible coefficient of performance is related to the temperatures of the hot and cold reservoirs:

$$\text{COP}_{max} = \frac{T_C}{T_H - T_C} \tag{11.15}$$

Theoretical maximum coefficient of performance of a heat pump used for cooling

Hot or cold lunch? Small coolers like this use Peltier devices that run off the 12 V electrical system of your car. They can transfer heat from the interior of the cooler to the outside, keeping food or drinks cool. But Peltier devices, like some other heat pumps, are reversible; switching the direction of current reverses the direction of heat transfer. A flick of a switch will cause heat to be transferred into the interior, keeping your lunch warm!

A refrigerator from 1934.

We can also use a heat pump for heating, moving heat from a cold reservoir to a hot reservoir to keep it warm. In that case, we define the coefficient of performance as

$$\text{COP} = \frac{\text{what you get}}{\text{what you had to pay}} = \frac{\text{energy added to the hot reservoir}}{\text{work required to perform the transfer}} = \frac{Q_H}{W_{in}}$$

In this case, the maximum possible coefficient of performance is

$$\text{COP}_{max} = \frac{T_H}{T_H - T_C} \tag{11.16}$$

Theoretical maximum coefficient of performance of a heat pump used for heating

In both cases, **a larger coefficient of performance means a more efficient heat pump.** Unlike the efficiency of a heat engine, which must be less than 1, the COP of a heat pump can be—and usually is—greater than 1. The following example shows that the COP can be quite high for typical temperatures.

EXAMPLE 11.13 Coefficient of performance of a refrigerator

The inside of your refrigerator is approximately 0°C. Heat from the inside of your refrigerator is deposited into the air in your kitchen, which has a temperature of approximately 20°C. At these operating temperatures, what is the maximum possible coefficient of performance of your refrigerator?

PREPARE The temperatures of the hot side and the cold side must be expressed in kelvin:

$$T_H = 20°\text{C} = 293 \text{ K} \qquad T_C = 0°\text{C} = 273 \text{ K}$$

SOLVE We use Equation 11.15 to compute the maximum coefficient of performance:

$$\text{COP}_{max} = \frac{T_C}{T_H - T_C} = \frac{273 \text{ K}}{293 \text{ K} - 273 \text{ K}} = 13.6$$

ASSESS A coefficient of performance of 13.6 means that we pump 13.6 J of heat for an energy cost of 1 J. Due to practical limitations, the coefficient of performance of an actual refrigerator is typically ≈5. Other factors affect the overall efficiency of the appliance, including how well insulated it is.

CONCEPTUAL EXAMPLE 11.14 Keeping your cool?

It's a hot day, and your apartment is rather warm. Your roommate suggests cooling off the apartment by keeping the door of the refrigerator open. Will this help the situation?

REASON It's time for a physics lesson for your roommate! What you want to do is remove heat from the room. Your refrigerator, a heat pump, is designed to transfer heat from its inside to the outside. If you leave the door open, all parts of the refrigerator are exposed to the room air. It simply transfers heat from one part of the room to another—it won't make the room cooler.

ASSESS An air conditioner is a heat pump too. It must have a hot side that is outside the house, so that there is a net transfer of heat from inside to outside.

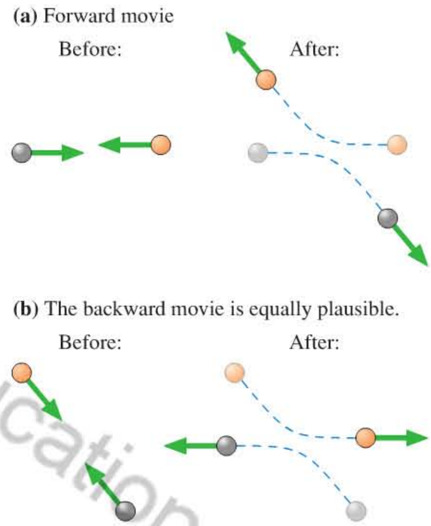
STOP TO THINK 11.6 Which of the following changes would allow your refrigerator to use less energy to run? (There may be more than one correct answer.)

A. Increasing the temperature inside the refrigerator
B. Increasing the temperature of the kitchen
C. Decreasing the temperature inside the refrigerator
D. Decreasing the temperature of the kitchen

11.8 Entropy and the Second Law of Thermodynamics

Throughout the chapter, we have noticed certain trends and certain limitations in energy transformations and transfers. Heat is transferred spontaneously from hot to cold, not from cold to hot. Once energy is transformed to thermal energy, it is (in some sense) "lost." The spontaneous transfer of heat from hot to cold is an example of an **irreversible** process, a process that can happen in only one direction. Why are some processes irreversible? The spontaneous transfer of heat from cold to hot would not violate any law of physics that we have seen to this point, but it is never observed. There must be another law of physics that prevents it. In this section we will explore the basis for this law, which we will call the **second law of thermodynamics.**

FIGURE 11.24 Molecular collisions are reversible.



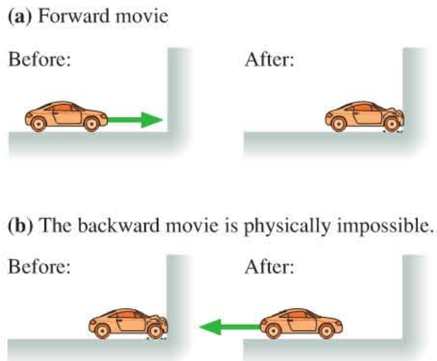
Reversible and Irreversible Processes

Stirring the cream in your coffee mixes the cream and coffee together. No amount of stirring ever unmixes them. If you watched a movie of someone stirring a cup of coffee and unmixing the cream, you'd be quite certain that the movie was running backward. In fact, a reasonable definition of an irreversible process is one for which a backward-running movie shows a physically impossible process.

At a microscopic level, collisions between molecules are completely reversible. In FIGURE 11.24 we see two possible movies of a collision between two gas molecules, one forward and the other backward. You can't tell by looking which is really going forward and which is being played backward. Nothing in either collision looks wrong, and no measurements you might make on either would reveal any violations of Newton's laws. Interactions at the molecular level are reversible processes.

At a macroscopic level, it's a different story. FIGURE 11.25 shows two possible movies of the collision of a car with a barrier. One movie is being run forward, the other backward. The backward movie of Figure 11.25b is obviously wrong. But what has been violated in the backward movie? To have the car return to its original shape and spring away from the wall would not violate any laws of physics we have so far discussed.

FIGURE 11.25 Macroscopic collisions are not reversible.



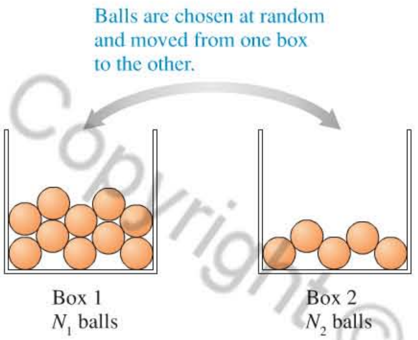
If microscopic motions are all reversible, how can macroscopic phenomena such as the car crash end up being irreversible? If reversible collisions can cause heat to be transferred from hot to cold, why do they never cause heat to be transferred from cold to hot? There must be something at work that can distinguish the past from the future.

Which Way to Equilibrium?

How do two systems initially at different temperatures "know" which way to go to reach equilibrium? Perhaps an analogy will help.

FIGURE 11.26 on the next page shows two boxes, numbered 1 and 2, containing identical balls. Box 1 starts with more balls than box 2, so $N_{1i} > N_{2i}$. Once every second, a ball in one of the two boxes is chosen at random and moved to the other box. This is a reversible process because a ball can move from box 2 to box 1 just as easily as from box 1 to box 2. What do you expect to see if you return several hours later?

FIGURE 11.26 Moving balls between boxes.



Because balls are chosen at random, and because $N_{1i} > N_{2i}$, it's initially more likely that a ball will move from box 1 to box 2 than from box 2 to box 1. Sometimes a ball will move "backward" from box 2 to box 1, but overall there's a net movement of balls from box 1 to box 2. The system will evolve until $N_1 \approx N_2$. We have reached a stable situation—equilibrium!—with an equal number of balls moving in both directions.

But couldn't it go the other way, with N_1 getting even larger while N_2 decreases? In principle, any arrangement of the balls is possible. But certain arrangements are more likely. Each ball is equally likely to be in either box. With four balls, the odds are 1 in 2^4, or 1 in 16, that, at a randomly chosen instant of time, you would find all the balls in box 1. Were you to do so, you wouldn't find that to be terribly surprising. But with 100 balls, the probability has dropped to about 1 in 1,000,000,000,000,000,000,000,000,000,000, or 1 in 10^{30}, that all of the balls will be in box 1.

Although each transfer is reversible, **the statistics of large numbers make it overwhelmingly likely that the system will evolve toward a state in which $N_1 \approx N_2$.** For the 10^{23} or so particles in a realistic system of atoms or molecules, we will never see a state that deviates appreciably from the equilibrium case.

So imagine: Is it possible that all of the air molecules in the room in which you are sitting could, in the next second, end up moving to one side of the room, leaving the other side empty? In principle it's possible, but this situation is so extraordinarily unlikely that we can, in good conscience, call it impossible.

A system reaches thermal equilibrium not by any plan or by outside intervention, but simply because **equilibrium is the most probable state in which to be.** The consequence of having a vast number of random events is that the system evolves in one direction, toward equilibrium, and not the other. Reversible microscopic events lead to irreversible macroscopic behavior because some macroscopic states are vastly more probable than others.

Order, Disorder, and Entropy

FIGURE 11.27 Ordered and disordered arrangements of atoms in a gas.

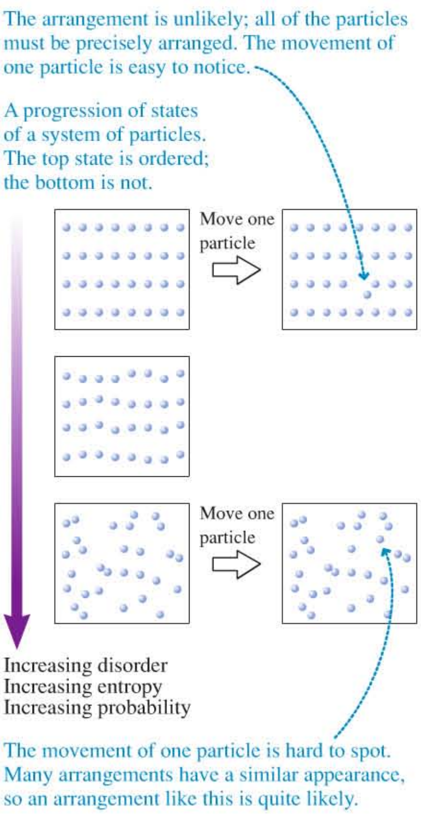


FIGURE 11.27 shows microscopic views of three containers of gas. The top diagram shows a group of atoms arranged in a regular pattern. This is a highly ordered and nonrandom system, with each atom's position precisely specified. Contrast this with the system on the bottom, in which there is no order at all. It is extremely improbable that the atoms in a container would *spontaneously* arrange themselves into the ordered pattern of the top picture; even a small change in the pattern is quite noticeable. By contrast, there are a vast number of arrangements like the one on the bottom that randomly fill the container. A small change in the pattern is hard to detect.

Scientists and engineers use the term **entropy** to quantify the probability that a certain state of a system will occur. The ordered arrangement of the top system, which has a very small probability of spontaneous occurrence, has a very low entropy. The entropy of the randomly filled container is high; it has a large probability of occurrence. The entropy in Figure 11.27 increases as you move from the ordered system on the top to the disordered system on the bottom.

Suppose you ordered the atoms in Figure 11.27 as they are in the top diagram, and then you let them go. After some time, you would expect their arrangement to appear as in the bottom diagram. In fact, the series of diagrams from top to bottom may be thought of as a series of movie frames showing the evolution of the positions of the atoms after they are released, moving toward a state of higher entropy. You would correctly expect this process to be irreversible—the particles will never spontaneously recreate their initial ordered state.

Two thermally interacting systems with different temperatures have a low entropy. These systems are ordered: The faster atoms are on one side of the barrier, the slower atoms on the other. The most random possible distribution of energy, and hence the least ordered system, corresponds to the situation where the two systems are in thermal equilibrium with equal temperatures. **Entropy increases as two systems with initially different temperatures move toward thermal equilibrium.**

TRY IT YOURSELF

▶ Typing Shakespeare Make a new document in your word processor. Close your eyes and type randomly for a while. Now open your eyes. Did you type any recognizable words? There is a chance that you did, but you probably didn't. One thousand chimps in a room, typing away randomly, *could* type the works of Shakespeare. Molecular collisions *could* transfer energy from a cold object to a hot object. But, the probability is so tiny that the outcome is never seen in the real world.

The fact that macroscopic systems evolve irreversibly toward equilibrium is a new law of physics, **the second law of thermodynamics:**

Second law of thermodynamics The entropy of an isolated system never decreases. The entropy either increases, until the system reaches equilibrium, or, if the system began in equilibrium, stays the same.

NOTE ▶ The qualifier "isolated" is crucial. We can order the system by reaching in from the outside, perhaps using little tweezers to place atoms in a lattice. Similarly, we can transfer heat from cold to hot by using a refrigerator. The second law is about what a system can or cannot do *spontaneously,* on its own, without outside intervention. ◀

The second law of thermodynamics tells us that an isolated system evolves such that:

- Order turns into disorder and randomness.
- Information is lost rather than gained.
- The system "runs down" as other forms of energy are transformed into thermal energy.

An isolated system never spontaneously generates order out of randomness. It is not that the system "knows" about order or randomness, but rather that there are vastly more states corresponding to randomness than there are corresponding to order. As collisions occur at the microscopic level, the laws of probability dictate that the system will, on average, move inexorably toward the most probable and thus most random macroscopic state.

Entropy and Thermal Energy

Suppose we have a very cold, moving baseball, with essentially no thermal energy, and a stationary helium balloon at room temperature, as shown in **FIGURE 11.28**. The atoms in the baseball and the atoms in the balloon are all moving, but there is a big difference in their motions. The atoms in the baseball are all moving in the same direction at the same speed, but the atoms in the balloon are moving in random directions. The ordered, organized motion of the baseball has low entropy, while the disorganized, random motion of the gas atoms—what we have called thermal energy—has high entropy. You can see that a conversion of macroscopic kinetic energy into thermal energy means an increase in entropy. We saw, in Section 11.5, that the conversion of other forms of energy into thermal energy was irreversible. Now, we can explain why: **When another form of energy is converted into thermal energy, there is an increase in entropy.**

This is why converting thermal energy into other forms of energy can't be done with 100% efficiency. We have redrawn our schematic diagram of a heat engine in **FIGURE 11.29** on the next page, adding arrows representing entropy. Now we can see why heat must be exhausted into the cold reservoir: We need to get rid of entropy! Efficiencies or coefficients of performance higher than those in Equations 11.14, 11.15, and 11.16 would reduce entropy and thus violate the second law of thermodynamics.

FIGURE 11.28 Kinetic energy and thermal energy compared.

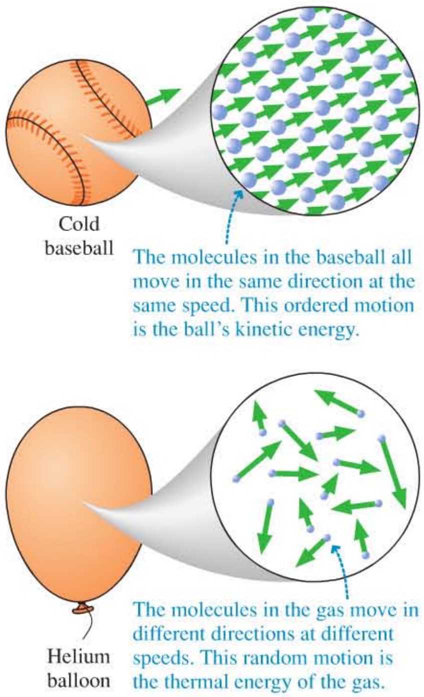
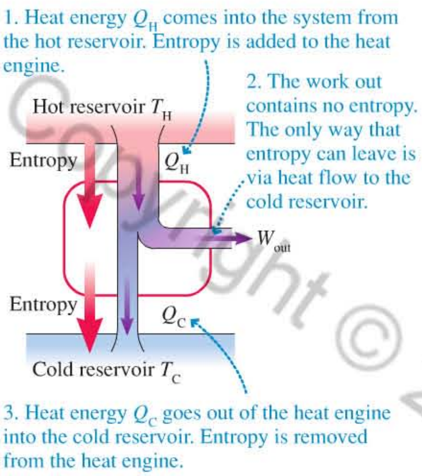


FIGURE 11.29 A diagram of a heat engine, with entropy changes noted.



Of all of the forms of energy we've seen, only thermal energy has entropy; the others are all ordered. As long as thermal energy isn't involved, you can freely and reversibly convert between different forms of energy, as we've seen. When you toss a ball into the air, the kinetic energy is converted to potential energy as the ball rises, then back to kinetic energy as the ball falls. But when some form of energy is transformed into thermal energy, entropy has increased, and the second law of thermodynamics tells us that this change is irreversible. When you drop a ball on the floor, it bounces several times and then stops moving. The kinetic energy of the ball in motion has become thermal energy of the slightly warmer ball at rest. This process won't reverse—the ball won't suddenly cool and jump into the air! This process would conserve energy, but it would reduce entropy, and so would violate this new law of physics.

This property of thermal energy has consequences for the efficiencies of heat engines, as we've seen, but it has consequences for the efficiencies of other devices as well.

CONCEPTUAL EXAMPLE 11.15 Efficiency of hybrid vehicles

Hybrid vehicles are powered by a gasoline engine paired with an electric motor and batteries. They get much better mileage in the stop and go of city driving than conventional vehicles do. When you brake to a stop in a conventional car, friction converts the kinetic energy of the car's motion into thermal energy in the brakes. In a typical hybrid car, some of this energy is converted into chemical energy in a battery. Explain how this makes a hybrid vehicle more efficient.

REASON When energy is transformed into thermal energy, the increase in entropy makes this change irreversible. When you brake a conventional car, the kinetic energy is transformed into the thermal energy of hot brakes and is lost to your use. In a hybrid vehicle, the kinetic energy is converted into chemical energy in a battery. This change is reversible; when the car starts again, the energy can be transformed back into kinetic energy.

ASSESS Whenever energy is converted to thermal energy, it is in some sense "lost," which reduces efficiency. The hybrid vehicle avoids this transformation, and so is more efficient.

STOP TO THINK 11.7 Which of the following processes does not involve a change in entropy?

A. An electric heater raises the temperature of a cup of water by 20°C.
B. A ball rolls up a ramp, decreasing in speed as it rolls higher.
C. A basketball is dropped from 2 m and bounces until it comes to rest.
D. The sun shines on a black surface and warms it.

11.9 Systems, Energy, and Entropy

Over the past two chapters, we have learned a good deal about energy and how it is used. By introducing the concept of entropy, we were able to see how limits on our ability to use energy come about. We will close this chapter, and this part of the book, by considering a few final questions that bring these different pieces together.

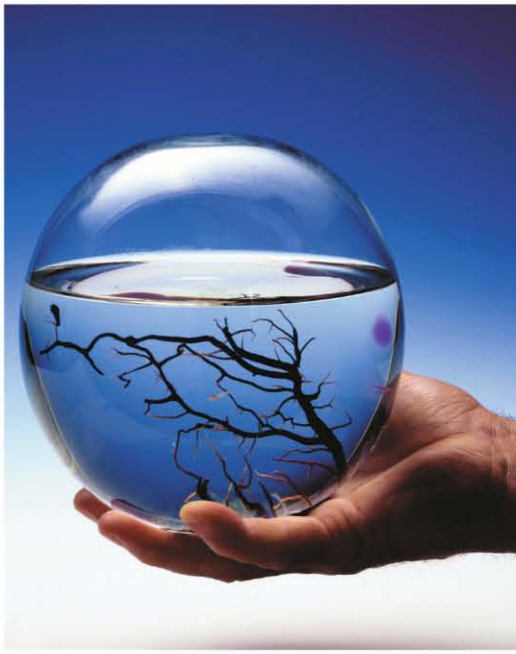
The Conservation of Energy and Energy Conservation

We have all heard for many years that it is important to "conserve energy." We are asked to turn off lights when we leave rooms, to drive our cars less, to turn down our thermostats. But this brings up an interesting question: If we have a law of conservation of energy, which states that energy can't be created or destroyed, what do we really mean by "conserving energy"? If energy can't be created or destroyed, how can there be an "energy crisis"?

We started this chapter looking at energy transformations. We saw that whenever energy is transformed, some of it is "lost." And now we know what this means: The energy isn't really lost, but it is converted into thermal energy. This change is irreversible; thermal energy can't be efficiently converted back into other forms of energy.

And that's the problem. We aren't, as a society or as a planet, running out of energy. We can't! What we can run out of is high-quality sources of energy. Oil is a good example. A gallon of gasoline contains a great deal of chemical energy. It is a liquid, so is easily transported, and it is easily burned in a host of devices to generate heat, electricity, or motion. When you burn gasoline in your car, you don't use up its energy—you simply convert its chemical energy into thermal energy. As you do this, you decrease the amount of high-quality chemical energy in the world and increase the supply of thermal energy. The amount of energy in the world is still the same; it's just in a less useful form.

Perhaps the best way to "conserve energy" is to concentrate on efficiency, to reduce "what you had to pay." More efficient lightbulbs, more efficient cars—all of these use less energy to produce the same final result.



Sealed, but not isolated BIO This glass container is a completely sealed system containing living organisms, shrimp and algae. But the organisms will live and grow for many years. The reason this is possible is that the glass sphere, though sealed, is not an *isolated* system. Energy can be transferred in and out as light and heat. If the container were placed in a darkened room, the organisms would quickly perish.

Entropy and Life

The second law of thermodynamics predicts that systems will "run down," that ordered states will evolve toward disorder and randomness. But living organisms seem to violate this rule:

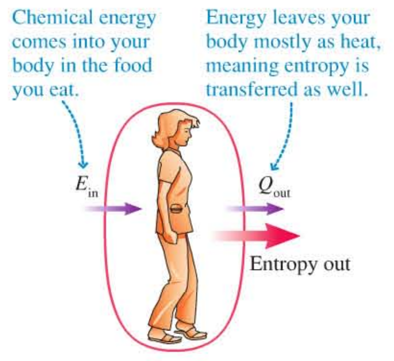
- Plants grow from simple seeds to complex entities.
- Single-celled fertilized eggs grow into complex adult organisms.
- Over the last billion years or so, life has evolved from simple unicellular organisms to very complex forms.

How can this be?

There is an important qualification in the second law of thermodynamics: It applies only to isolated systems, systems that do not exchange energy with their environment. The situation is entirely different if energy is transferred into or out of the system.

Your body is not an isolated system. Every day, you take in chemical energy in the food you eat. As you use this energy, most of it ends up as thermal energy that you exhaust as heat into the environment, taking away the entropy created by natural processes in your body. An energy and entropy diagram of this situation is given in **FIGURE 11.30**. Each second, as you sit quietly and read this book, your body is using 100 J of chemical energy and exhausting 100 J of thermal energy to the environment. The entropy of your body is staying approximately constant, but the entropy of the environment is increasing due to the thermal energy from your body. To grow and develop, organisms must take in high-quality forms of energy and exhaust thermal energy. This continuous exchange of energy (and entropy) with the environment makes your life—and all life—possible without violating any laws of physics.

FIGURE 11.30 Thermodynamic view of the body.



INTEGRATED EXAMPLE 11.16 Efficiency of an automobile

In the absence of external forces, Newton's first law tells us that a car would continue at a constant speed once it was moving. So why can't you get your car up to speed and then take your foot off the gas? Because there *are* external forces. At highway speeds, nearly all of the force opposing your car's motion is the drag force of the air. (Rolling friction is much smaller than drag at highway speeds, so we'll ignore it.) **FIGURE 11.31** shows that as your car moves through the air it pushes the air aside. Doing so takes energy, and your car's engine must keep running to replace this lost energy.

FIGURE 11.31 A wind-tunnel test shows airflow around a car.

Sports cars, with their aerodynamic shape, are often those that lose the least energy to drag. A typical sports car has a 350 hp engine, a drag coefficient of 0.30, and a low profile with the area of the front of the car being a modest 1.8 m^2. Such a car gets about 25 miles per gallon of gasoline at a highway speed of 30 m/s (just over 65 mph).

Suppose this car is driven 25 miles at 30 m/s. It will consume 1 gallon of gasoline, containing 1.4×10^8 J of chemical energy. What is the car's efficiency for this trip?

PREPARE We can use what we learned about the drag force in Chapter 5 to compute the amount of energy needed to move the car forward through the air. We can use this value in Equation 11.2 as what you get; it is the minimum amount of energy that *could* be used to move the car forward at this speed. We can then use the mileage data to calculate what you had to pay; this is the energy actually used by the car's engine. Once we have these two pieces of information, we can calculate efficiency.

SOLVE Heat and work both play a role in this process, so we need to use our most general equation about energy and energy conservation, Equation 11.11:

$$\Delta K + \Delta U + \Delta E_{\text{th}} + \Delta E_{\text{chem}} + \cdots = W + Q$$

The car is moving at a constant speed, so its kinetic energy is not changing. The road is assumed to be level, so there is no change in potential energy. Once the car is warmed up, its temperature will be constant, so its thermal energy isn't changing. Equation 11.11 reduces to

$$\Delta E_{\text{chem}} = W + Q$$

The car's engine is using chemical energy as it burns fuel. The engine transforms some of this energy to work—the propulsion force pushing the car forward against the opposing force of air drag. But the engine also transforms much of the chemical energy to "waste heat" that is transferred to the environment through the radiator and the exhaust. ΔE_{chem} is negative because the amount of energy stored in the gas tank is decreasing as gasoline is burned. W and Q are also negative, according to the sign convention of Figure 11.16, because these energies are being transferred *out* of the system to the environment.

In Chapter 5 we saw that the drag force on an object moving at a speed v is given by

$$\vec{D} = \left(\tfrac{1}{2}C_{\text{D}}\rho A v^2, \text{ direction opposite the motion}\right)$$

where C_{D} is the drag coefficient, ρ is the density of air (approximately 1.2 kg/m^3), and A is the area of the front of the car. The drag force on the car moving at 30 m/s is

$$D = \tfrac{1}{2}(0.30)(1.2\ \text{kg/m}^3)(1.8\ \text{m}^2)(30\ \text{m/s})^2 = 290\ \text{N}$$

The drag force opposes the motion. To move the car forward at constant speed, and thus with no *net* force, requires a propulsion force $F = 290$ N in the forward direction. In Chapter 10 we found that the power—the rate of energy expenditure—to move an object at speed v with force F is $P = Fv$. Thus the power the car must supply—at the wheels—to keep the car moving down the highway is

$$P = Fv = (290\ \text{N})(30\ \text{m/s}) = 8700\ \text{W}$$

It's interesting to compare this to the engine power by converting to horsepower:

$$P = 8700\ \text{W}\left(\frac{1\ \text{hp}}{746\ \text{W}}\right) = 12\ \text{hp}$$

Only a small fraction of the engine's 350 horsepower is needed to keep the car moving at highway speeds.

The distance traveled is

$$\Delta x = 25\ \text{mi} \times \frac{1.6\ \text{km}}{1\ \text{mi}} \times \frac{1000\ \text{m}}{1\ \text{km}} = 40{,}000\ \text{m}$$

and the time required to travel it is $\Delta t = (40{,}000\ \text{m})/(30\ \text{m/s}) = 1333$ s. Thus the minimum energy needed to travel 40 km—the energy needed simply to push the air aside—is

$$E_{\text{min}} = P\Delta t = (8700\ \text{W})(1333\ \text{s}) = 1.2 \times 10^7\ \text{J}$$

In traveling this distance, the car uses 1 gallon of gas, with 1.4×10^8 J of chemical energy. This is, quite literally, what you had to pay to drive this distance. Thus the car's efficiency is

$$e = \frac{\text{what you get}}{\text{what you had to pay}} = \frac{1.2 \times 10^7\ \text{J}}{1.4 \times 10^8\ \text{J}} = 0.086 = 8.6\%$$

ASSESS The efficiency of the car is quite low, even compared to other engines that we've seen. Nonetheless, our calculation agrees reasonably well with actual measurements. Gasoline-powered vehicles simply are inefficient, which is one factor favoring more efficient alternative vehicles. Smaller mass and better aerodynamic design would improve the efficiency of vehicles, but a large part of the inefficiency of a gasoline-powered vehicle is inherent in the thermodynamics of the engine itself and in the complex drive train needed to transfer the engine's power to the wheels.

SUMMARY

The goals of Chapter 11 have been to learn more about energy transformations and transfers, the laws of thermodynamics, and theoretical and practical limitations on energy use.

GENERAL PRINCIPLES

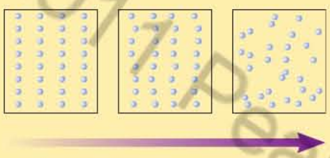
Energy and Efficiency

When energy is transformed from one form into another, some may be "lost," usually to thermal energy, due to practical or theoretical constraints. This limits the efficiency of processes. We define **efficiency** as

$$e = \frac{\text{what you get}}{\text{what you had to pay}}$$

Entropy and Irreversibility

Systems move toward more probable states. These states have higher **entropy**—more disorder. This change is irreversible. Changing other forms of energy to thermal energy is irreversible.



The Laws of Thermodynamics

The **first law of thermodynamics** is a statement of conservation of energy for systems in which only thermal energy changes:

$$\Delta E_{th} = W + Q$$

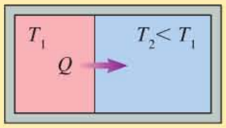
The **second law of thermodynamics** specifies the way that isolated systems can evolve:

The entropy of an isolated system always increases.

This law has practical consequences:

- Heat energy spontaneously flows only from hot to cold.
- A transformation of energy into thermal energy is irreversible.
- No heat engine can be 100% efficient.

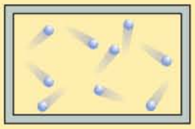
Heat is energy transferred between two objects at different temperatures. Energy will be transferred until thermal equilibrium is reached.



IMPORTANT CONCEPTS

Thermal energy

- For a gas, the thermal energy is the **total kinetic energy** of motion of the atoms.
- Thermal energy is random kinetic energy and so has entropy.



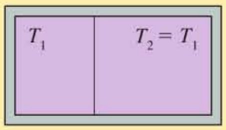
$$E_{th} = NK_{avg} = \frac{3}{2}Nk_B T$$

Temperature

- For a gas, temperature is proportional to the **average kinetic energy** of the motion of the atoms.

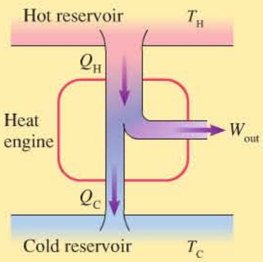
$$T = \frac{2}{3}\frac{K_{avg}}{k_B}$$

- Two systems are in **thermal equilibrium** if they are at the same temperature. No heat energy is transferred at thermal equilibrium.



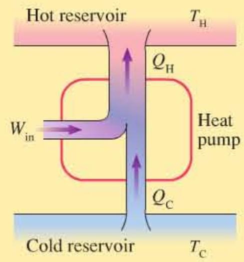
A **heat engine** converts thermal energy from a hot reservoir into useful work. Some heat is exhausted into a cold reservoir, limiting efficiency.

$$e_{max} = 1 - \frac{T_C}{T_H}$$



A **heat pump** uses an energy input to transfer heat from a cold side to a hot side. The **coefficient of performance** is analogous to efficiency. For cooling, it is:

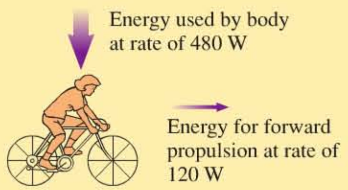
$$\text{COP}_{max} = \frac{T_C}{T_H - T_C}$$



APPLICATIONS

Efficiencies

Energy in the body
Cells in the body metabolize chemical energy in food. Efficiency for most actions is about 25%.



Power plants
A typical power plant converts about 1/3 of the energy input into useful work. The rest is exhausted as waste heat.

Temperature scales

Zero on the **Kelvin temperature scale** is the temperature at which the kinetic energy of atoms is zero. This is **absolute zero.** The conversion from °C to K is

$$T(\text{K}) = T(°\text{C}) + 273$$

▶ All temperatures in equations must be in kelvin. ◀

For homework assigned on MasteringPhysics, go to www.masteringphysics.com

Problems labeled INT integrate significant material from earlier chapters; BIO are of biological or medical interest.

Problem difficulty is labeled as | (straightforward) to ||||| (challenging).

QUESTIONS

Conceptual Questions

1. Rub your hands together vigorously. What happens? Discuss the energy transfers and transformations that take place.
2. Write a few sentences describing the energy transformations that occur from the time moving water enters a hydroelectric plant until you see some water being pumped out of a nozzle in a public fountain. Use the "Energy transformations" table on page 324 as an example.
3. BIO Describe the energy transfers and transformations that occur from the time you sit down to breakfast until you've completed a fast bicycle ride.
4. BIO According to Table 11.4, cycling at 15 km/h requires less metabolic energy than running at 15 km/h. Suggest reasons why this is the case.
5. You're stranded on a remote desert island with only a chicken, a bag of corn, and a shade tree. To survive as long as possible in hopes of being rescued, should you eat the chicken at once and then the corn? Or eat the corn, feeding enough to the chicken to keep it alive, and then eat the chicken when the corn is gone? Or are your survival chances the same either way? Explain.
6. For most automobiles, the number of miles per gallon decreases as highway speed increases. Fuel economy drops as speeds increase from 55 to 65 mph, then decreases further as speeds increase to 75 mph. Explain why this is the case.
7. When the space shuttle returns to earth, its surfaces get very hot as it passes through the atmosphere at high speed.
   a. Has the space shuttle been heated? If so, what was the source of the heat? If not, why is it hot?
   b. Energy must be conserved. What happens to the space shuttle's initial kinetic energy?
8. One end of a short aluminum rod is in a campfire and the other end is in a block of ice, as shown in Figure Q11.8. If 100 J of energy are transferred from the fire to the rod, and if the temperature at every point in the rod has reached a steady value, how much energy goes from the rod into the ice?

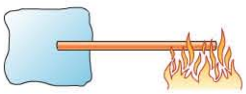


FIGURE Q11.8

9. Two blocks of copper, one of mass 1 kg and the second of mass 3 kg, are at the same temperature. Which block has more thermal energy? If the blocks are placed in thermal contact, will the thermal energy of the blocks change? If so, how?
10. If the temperature T of an ideal gas doubles, by what factor does the average kinetic energy of the atoms change?
11. A bottle of helium gas and a bottle of argon gas contain equal numbers of atoms at the same temperature. Which bottle, if either, has the greater total thermal energy?

For Questions 12 through 17, give a specific example of a process that has the energy changes and transfers described. (For example, if the question states "$\Delta E_{th} > 0$, $W = 0$," you are to describe a process that has an increase in thermal energy and no transfer of energy by work. You could write "Heating a pan of water on the stove.")

12. $\Delta E_{th} < 0$, $W = 0$
13. $\Delta E_{th} > 0$, $Q = 0$
14. $\Delta E_{th} < 0$, $Q = 0$
15. $\Delta E_{th} > 0$, $W \neq 0$, $Q \neq 0$
16. $\Delta E_{th} < 0$, $W \neq 0$, $Q \neq 0$
17. $\Delta E_{th} = 0$, $W \neq 0$, $Q \neq 0$
18. A fire piston—an impressive physics demonstration—ignites a fire without matches. The operation is shown in Figure Q11.18. A wad of cotton is placed at the bottom of a sealed syringe with a tight-fitting plunger. When the plunger is rapidly depressed, the air temperature in the syringe rises enough to ignite the cotton. Explain why the air temperature rises, and why the plunger must be pushed in very quickly.

FIGURE Q11.18

19. In a gasoline engine, fuel vapors are ignited by a spark. In a diesel engine, a fuel–air mixture is drawn in, then rapidly compressed to as little as 1/20 the original volume, in the process increasing the temperature enough to ignite the fuel–air mixture. Explain why the temperature rises during the compression.
20. A drop of green ink falls into a beaker of clear water. First *describe* what happens. Then *explain* the outcome in terms of entropy.
21. If you hold a rubber band loosely between two fingers and then stretch it, you can tell by touching it to the sensitive skin of your forehead that stretching the rubber band has increased its temperature. If you then let the rubber band rest against your forehead, it soon returns to its original temperature. What are the signs of W and Q for the entire process?
22. In areas in which the air temperature drops very low in the winter, the exterior unit of a heat pump designed for heating is sometimes buried underground in order to use the earth as a thermal reservoir. Why is it worthwhile to bury the heat exchanger, even if the underground unit costs more to purchase and install than one above ground?
23. Assuming improved materials and better processes, can engineers ever design a heat engine that exceeds the maximum efficiency indicated by Equation 11.14? If not, why not?
24. Electric vehicles increase speed by using an electric motor that draws energy from a battery. When the vehicle slows, the motor runs as a generator, recharging the battery. Explain why this means that an electric vehicle can be more efficient than a gasoline-fueled vehicle.
25. When the sun's light hits the earth, the temperature rises. Is there an entropy change to accompany this transformation? Explain.

26. When you put an ice cube tray filled with liquid water in your freezer, the water eventually becomes solid ice. The solid is more ordered than the liquid—it has less entropy. Explain how this transformation is possible without violating the second law of thermodynamics.
27. A company markets an electric heater that is described as 100% efficient at converting electric energy to thermal energy. Does this violate the second law of thermodynamics?

Multiple-Choice Questions

28. ||| A person is walking on level ground at constant speed. What energy transformation is taking place?
    A. Chemical energy is being transformed to thermal energy.
    B. Chemical energy is being transformed to kinetic energy.
    C. Chemical energy is being transformed to kinetic energy and thermal energy.
    D. Chemical energy and thermal energy are being transformed to kinetic energy.
29. | A person walks 1 km, turns around, and runs back to where he started. Compare the energy used and the power during the two segments.
    A. The energy used and the power are the same for both.
    B. The energy used while walking is greater, the power while running is greater.
    C. The energy used while running is greater, the power while running is greater.
    D. The energy used is the same for both segments, the power while running is greater.
30. | The temperature of the air in a basketball increases as it is pumped up. This means that
    A. The total kinetic energy of the air is increasing and the average kinetic energy of the molecules is decreasing.
    B. The total kinetic energy of the air is increasing and the average kinetic energy of the molecules is increasing.
    C. The total kinetic energy of the air is decreasing and the average kinetic energy of the molecules is decreasing.
    D. The total kinetic energy of the air is decreasing and the average kinetic energy of the molecules is increasing.
31. | The thermal energy of a container of helium gas is halved. What happens to the temperature, in kelvin?
    A. It decreases to one-fourth its initial value.
    B. It decreases to one-half its initial value.
    C. It stays the same.
    D. It increases to twice its initial value.
32. | An inventor approaches you with a device that he claims will take 100 J of thermal energy input and produce 200 J of electricity. You decide not to invest your money because this device would violate
    A. The first law of thermodynamics.
    B. The second law of thermodynamics.
    C. Both the first and second laws of thermodynamics.
33. | While keeping your food cold, your refrigerator transfers energy from the inside to the surroundings. Thus thermal energy goes from a colder object to a warmer one. What can you say about this?
    A. It is a violation of the second law of thermodynamics.
    B. It is not a violation of the second law of thermodynamics because refrigerators can have efficiency of 100%.
    C. It is not a violation of the second law of thermodynamics because the second law doesn't apply to refrigerators.
    D. The second law of thermodynamics applies in this situation, but it is not violated because the energy did not spontaneously go from cold to hot.
34. || An electric power plant uses energy from burning coal to generate steam at 450°C. The plant is cooled by 20°C water from a nearby river. If burning coal provides 100 MJ of heat, what is the theoretical minimum amount of heat that must be transferred to the river during the conversion of heat to electric energy?
    A. 100 MJ B. 90 MJ
    C. 60 MJ D. 40 MJ
35. || A refrigerator's freezer compartment is set at –10°C; the kitchen is 24°C. What is the theoretical minimum amount of electric energy necessary to pump 1.0 J of energy out of the freezer compartment?
    A. 0.89 J B. 0.87 J
    C. 0.13 J D. 0.11 J

PROBLEMS

Section 11.1 Transforming Energy

1. || A 10% efficient engine accelerates a 1500 kg car from rest to 15 m/s. How much energy is transferred to the engine by burning gasoline?
2. || A 60% efficient device uses chemical energy to generate 600 J of electric energy.
    a. How much chemical energy is used?
    b. A second device uses twice as much chemical energy to generate half as much electric energy. What is its efficiency?
3. | A typical photovoltaic cell delivers 4.0×10^{-3} W of electric energy when illuminated with 1.2×10^{-1} W of light energy. What is the efficiency of the cell?
4. || An individual white LED (light-emitting diode) has an efficiency of 20% and uses 1.0 W of electric power. How many LEDs must be combined into one light source to give a total of 1.6 W of visible-light output (comparable to the light output of a 40 W incandescent bulb)? What total power is necessary to run this LED light source?

Section 11.2 Energy in the Body: Energy Inputs

5. || A fast-food hamburger (with cheese and bacon) contains BIO 1000 Calories. What is the burger's energy in joules?
6. ||| In an average human, basic life processes require energy to BIO be supplied at a steady rate of 100 W. What daily energy intake, in Calories, is required to maintain these basic processes?
7. | An "energy bar" contains 6.0 g of fat. How much energy is BIO this in joules? In calories? In Calories?
8. | An "energy bar" contains 22 g of carbohydrates. How much BIO energy is this in joules? In calories? In Calories?

Section 11.3 Energy in the Body: Energy Outputs

9. ||| An "energy bar" contains 22 g of carbohydrates. If the energy BIO bar was his only fuel, how far could a 68 kg person walk at 5.0 km/h?
10. ||| Suppose your body was able to use the chemical energy in BIO gasoline. How far could you pedal a bicycle at 15 km/h on the energy in 1 gal of gas? (1 gal of gas has a mass of 3.2 kg.)

11. The label on a candy bar says 400 Calories. Assuming a typical efficiency for energy use by the body, if a 60 kg person were to use the energy in this candy bar to climb stairs, how high could she go?
BIO

12. A weightlifter curls a 30 kg bar, raising it each time a distance of 0.60 m. How many times must he repeat this exercise to burn off the energy in one slice of pizza?
BIO

13. A weightlifter works out at the gym each day. Part of her routine is to lie on her back and lift a 40 kg barbell straight up from chest height to full arm extension, a distance of 0.50 m.
BIO
   a. How much work does the weightlifter do to lift the barbell one time?
   b. If the weightlifter does 20 repetitions a day, what total energy does she expend on lifting, assuming a typical efficiency for energy use by the body.
   c. How many 400 Calorie donuts can she eat a day to supply that energy?

Section 11.4 Thermal Energy and Temperature

14. The planet Mercury's surface temperature varies from 700 K during the day to 90 K at night. What are these values in °C and °F?
15. An ideal gas is at 20°C. If we double the average kinetic energy of the gas atoms, what is the new temperature in °C?
16. An ideal gas is at 20°C. The gas is cooled, reducing the thermal energy by 10%. What is the new temperature in °C?
17. An ideal gas at 0°C consists of 1.0×10^{23} atoms. 10 J of thermal energy are added to the gas. What is the new temperature in °C?
18. An ideal gas at 20°C consists of 2.2×10^{22} atoms. 4.3 J of thermal energy are removed from the gas. What is the new temperature in °C?

Section 11.5 Heat and the First Law of Thermodynamics

19. 500 J of work are done on a system in a process that decreases the system's thermal energy by 200 J. How much energy is transferred to or from the system as heat?
20. 600 J of heat energy are transferred to a system that does 400 J of work. By how much does the system's thermal energy change?
21. 300 J of energy are transferred to a system in the form of heat while the thermal energy increases by 150 J. How much work is done on or by the system?
22. 10 J of heat are removed from a gas sample while it is being compressed by a piston that does 20 J of work. What is the change in the thermal energy of the gas? Does the temperature of the gas increase or decrease?

Section 11.6 Heat Engines

23. A heat engine extracts 55 kJ from the hot reservoir and exhausts 40 kJ into the cold reservoir. What are (a) the work done and (b) the efficiency?
24. A heat engine does 20 J of work while exhausting 30 J of waste heat. What is the engine's efficiency?
25. A heat engine does 200 J of work while exhausting 600 J of heat to the cold reservoir. What is the engine's efficiency?
26. A heat engine with an efficiency of 40% does 100 J of work. How much heat is (a) extracted from the hot reservoir and (b) exhausted into the cold reservoir?
27. a. At what cold-reservoir temperature (in °C) would an engine operating at maximum theoretical efficiency with a hot-reservoir temperature of 427°C have an efficiency of 60%?
   b. If another engine, operating at maximum theoretical efficiency with a hot-reservoir temperature of 400°C, has the same efficiency, what is its cold-reservoir temperature?
28. A heat engine operating between energy reservoirs at 20°C and 600°C has 30% of the maximum possible efficiency. How much energy does this engine extract from the hot reservoir to do 1000 J of work?
29. A newly proposed device for generating electricity from the sun is a heat engine in which the hot reservoir is created by focusing sunlight on a small spot on one side of the engine. The cold reservoir is ambient air at 20°C. The designer claims that the efficiency will be 60%. What minimum hot-reservoir temperature, in °C, would be required to produce this efficiency?
30. Converting sunlight to electricity with solar cells has an efficiency of ≈15%. It's possible to achieve a higher efficiency (though currently at higher cost) by using concentrated sunlight as the hot reservoir of a heat engine. Each dish in Figure P11.30 concentrates sunlight on one side of a heat engine, producing a hot-reservoir temperature of 650°C. The cold reservoir, ambient air, is approximately 30°C. The actual working efficiency of this device is ≈30%. What is the theoretical maximum efficiency?

FIGURE P11.30

Section 11.7 Heat Pumps

31. A refrigerator takes in 20 J of work and exhausts 50 J of heat. What is the refrigerator's coefficient of performance?
32. Air conditioners are rated by their coefficient of performance at 80°F inside temperature and 95°F outside temperature. An efficient but realistic air conditioner has a coefficient of performance of 3.2. What is the maximum possible coefficient of performance?
33. 50 J of work are done on a refrigerator with a coefficient of performance of 4.0. How much heat is (a) extracted from the cold reservoir and (b) exhausted to the hot reservoir?
34. Find the maximum possible coefficient of performance for a heat pump used to heat a house in a northerly climate in winter. The inside is kept at 20°C while the outside is −20°C.

Section 11.8 Entropy and the Second Law of Thermodynamics

35. Which, if any, of the heat engines in Figure P11.35 below violate (a) the first law of thermodynamics or (b) the second law of thermodynamics? Explain.

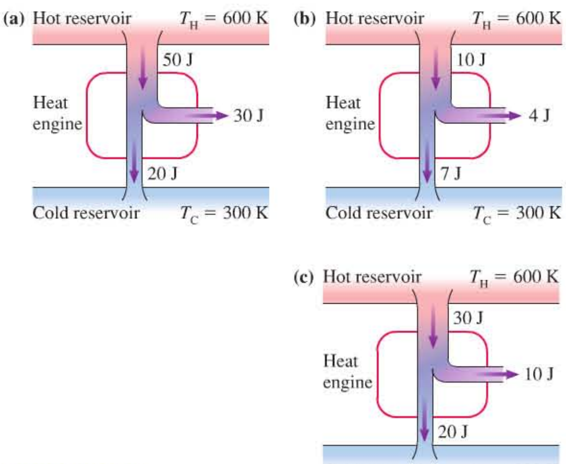


FIGURE P11.35

36. | Which, if any, of the refrigerators in Figure P11.36 below violate (a) the first law of thermodynamics or (b) the second law of thermodynamics? Explain.

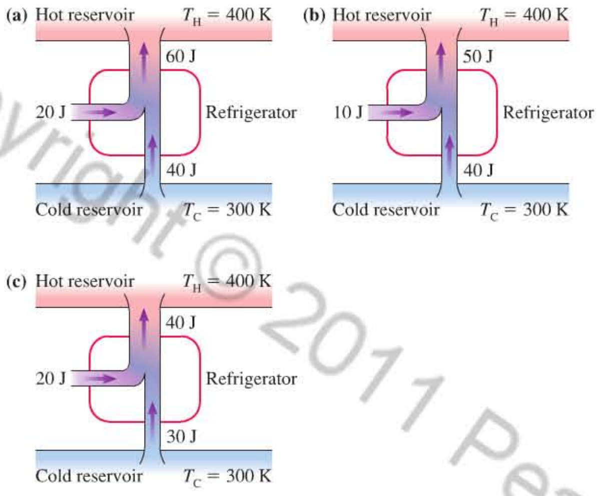


FIGURE P11.36

37. || Draw all possible distinct arrangements in which three balls (labeled A, B, C) are placed into two different boxes (1 and 2), as in Figure 11.26. If all arrangements are equally likely, what is the probability that all three will be in box 1?

General Problems

38. ||||| How many slices of pizza must you eat to walk for 1.0 h at a BIO speed of 5.0 km/h? (Assume your mass is 68 kg.)

39. || A 60 kg hiker climbs to the top of a 500-m-high hill. Ignoring the energy needed for horizontal motion and assuming a typical efficiency for energy use by the body, how many frozen burritos would be needed to fuel this climb?

40. ||||| For how long would a 68 kg athlete have to swim at a fast BIO crawl to use all the energy available in a typical fast-food meal of burger, fries, and a drink?

41. || BIO
   a. How much metabolic energy is required for a 68 kg runner to run at a speed of 15 km/h for 20 min?
   b. How much metabolic energy is required for this runner to walk at a speed of 5.0 km/h for 60 min? Compare your result to your answer to part a.
   c. Compare your results of parts a and b to the result of Example 11.4. Of these three modes of human motion, which is the most efficient?

42. | To a good approximation, the only external force that does BIO work on a cyclist moving on level ground is the force of air resistance. Suppose a cyclist is traveling at 15 km/h on level ground. Assume he is using 480 W of metabolic power.
   a. Estimate the amount of power he uses for forward motion.
   b. How much force must he exert to overcome the force of air resistance?

43. | The winning time for the 2005 annual race up 86 floors of the BIO Empire State Building was 10 min and 49 s. The winner's mass was 60 kg.
   a. If each floor was 3.7 m high, what was the winner's change in gravitational potential energy?
   b. If the efficiency in climbing stairs is 25%, what total energy did the winner expend during the race?
   c. How many food Calories did the winner "burn" in the race?
   d. Of those Calories, how many were converted to thermal energy?
   e. What was the winner's metabolic power in watts during the race up the stairs?

44. ||| Championship swimmers take about 22 s and about 30 arm BIO strokes to move through the water in a 50 m freestyle race. INT
   a. From Table 11.4, a swimmer's metabolic power is 800 W. If the efficiency for swimming is 25%, how much energy is expended moving through the water in a 50 m race?
   b. If half the energy is used in arm motion and half in leg motion, what is the energy expenditure per arm stroke?
   c. Model the swimmer's hand as a paddle. During one arm stroke, the paddle moves halfway around a 90-cm-radius circle. If all the swimmer's forward propulsion during an arm stroke comes from the hand pushing on the water and none from the arm (somewhat of an oversimplification), what is the average force of the hand on the water?

45. |||| A 68 kg hiker walks at 5.0 km/h up a 7% slope. What is the BIO necessary metabolic power? **Hint:** You can model her power needs as the sum of the power to walk on level ground plus the power needed to raise her body by the appropriate amount.

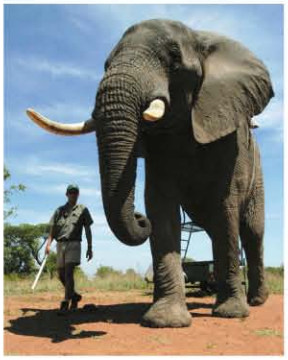
46. ||| A 70 kg student consumes 2500 Cal each day and stays the BIO same weight. One day, he eats 3500 Cal and, wanting to keep INT from gaining weight, decides to "work off" the excess by jumping up and down. With each jump, he accelerates to a speed of 3.3 m/s before leaving the ground. How many jumps must he make? Assume that the efficiency of the body in using energy is 25%

47. || To make your workouts more productive, you can get a gen- BIO erator that you drive with the rear wheel of your bicycle when it is mounted in a stand.
   a. Your laptop charger uses 75 W. What is your body's metabolic power use while running the generator to power your laptop charger, given the typical efficiency for such tasks? Assume 100% efficiency for the generator.
   b. Your laptop takes 1 hour to recharge. If you run the generator for 1 hour, how much energy does your body use? Express your result in joules and in Calories.

48. || The resistance of an exercise bike is often provided by a gen- BIO erator; that is, the energy that you expend is used to generate electric energy, which is then dissipated. Rather than dissipate the energy, it could be used for practical purposes.
   a. A typical person can maintain a steady energy expenditure of 400 W on a bicycle. Assuming a typical efficiency for the body and a generator that is 80% efficient, what useful electric power could you produce with a bicycle-powered generator?
   b. How many people would need to ride bicycle generators simultaneously to power a 400 W TV in the gym?

49. || Smaller mammals use proportionately more energy than BIO larger mammals; that is, it takes more energy per gram to power a mouse than a human. A typical mouse has a mass of 20 g and, at rest, needs to consume 3.0 Cal each day for basic body processes.
   a. If a 68 kg human used the same energy per kg of body mass as a mouse, how much energy would be needed each day?
   b. What resting power does this correspond to? How much greater is this than the resting power noted in the chapter?

50. BIO Larger animals use proportionately less energy than smaller animals; that is, it takes less energy per kg to power an elephant than to power a human. A 5000 kg African elephant requires about 70,000 Cal for basic needs for one day.



   a. If a 68 kg human required the same energy per kg of body mass as an elephant, how much energy would be required each day?
   b. What resting power does this correspond to? How much less is this than the resting power noted in the chapter?

51. BIO A large horse can perform work at a steady rate of about 1 horsepower, as you might expect.
   a. Assuming a 25% efficiency, how many Calories would a horse need to consume to work at 1.0 hp for 1.0 h?
   b. Dry hay contains about 10 MJ per kg. How many kilograms of hay would the horse need to eat to perform this work?

52. A college student is working on her physics homework in her dorm room. Her room contains a total of 6.0×10^{26} gas molecules. As she works, her body is converting chemical energy into thermal energy at a rate of 125 W. If her dorm room were an isolated system (dorm rooms can certainly feel like that) and if all of this thermal energy were transferred to the air in the room, by how much would the temperature increase in 10 min?

53. A container holding argon atoms changes temperature by 20°C when 30 J of heat are removed. How many atoms are in the container?

54. A heat engine with a high-temperature reservoir at 400 K has an efficiency of 0.20. What is the maximum possible temperature of the cold reservoir?

55. An engine does 10 J of work and exhausts 15 J of waste heat.
   a. What is the engine's efficiency?
   b. If the cold-reservoir temperature is 20°C, what is the minimum possible temperature in °C of the hot reservoir?

56. The heat exhausted to the cold reservoir of an engine operating at maximum theoretical efficiency is two-thirds the heat extracted from the hot reservoir. What is the temperature ratio T_C/T_H?

57. An engine operating at maximum theoretical efficiency whose cold-reservoir temperature is 7°C is 40% efficient. By how much should the temperature of the hot reservoir be increased to raise the efficiency to 60%?

58. Some heat engines can run on very small temperature differences. One manufacturer claims to have a very small heat engine that can run on the temperature difference between your hand and the air in the room. Estimate the theoretical maximum efficiency of this heat engine.

59. The coefficient of performance of a refrigerator is 5.0.
   a. If the compressor uses 10 J of energy, how much heat is exhausted to the hot reservoir?
   b. If the hot-reservoir temperature is 27°C, what is the lowest possible temperature in °C of the cold reservoir?

60. An engineer claims to have measured the characteristics of a heat engine that takes in 100 J of thermal energy and produces 50 J of useful work. Is this engine possible? If so, what is the smallest possible ratio of the temperatures (in kelvin) of the hot and cold reservoirs?

61. A 32% efficient electric power plant produces 900 MJ of electric energy per second and discharges waste heat into 20°C ocean water. Suppose the waste heat could be used to heat homes during the winter instead of being discharged into the ocean. A typical American house requires an average 20 kW for heating. How many homes could be heated with the waste heat of this one power plant?

62. A typical coal-fired power plant burns 300 metric tons of coal *every hour* to generate 2.7×10^6 MJ of electric energy. 1 metric ton = 1000 kg; 1 metric ton of coal has a volume of 1.5 m^3. The heat of combustion of coal is 28 MJ/kg. Assume that *all* heat is transferred from the fuel to the boiler and that *all* the work done in spinning the turbine is transformed into electric energy.
   a. Suppose the coal is piled up in a 10 m × 10 m room. How tall must the pile be to operate the plant for one day?
   b. What is the power plant's efficiency?

63. Each second, a nuclear power plant generates 2000 MJ of thermal energy from nuclear reactions in the reactor's core. This energy is used to boil water and produce high-pressure steam at 300°C. The steam spins a turbine, which produces 700 MJ of electric power, then the steam is condensed and the water is cooled to 30°C before starting the cycle again.
   a. What is the maximum possible efficiency of the plant?
   b. What is the plant's actual efficiency?

64. 250 students sit in an auditorium listening to a physics lecture. Because they are thinking hard, each is using 125 W of metabolic power, slightly more than they would use at rest. An air conditioner with a COP of 5.0 is being used to keep the room at a constant temperature. What minimum electric power must be used to operate the air conditioner?

65. INT Driving on asphalt roads entails very little rolling resistance, so most of the energy of the engine goes to overcoming air resistance. But driving slowly in dry sand is another story. If a 1500 kg car is driven in sand at 5.0 m/s, the coefficient of rolling friction is 0.06. In this case, nearly all of the energy that the car uses to move goes to overcoming rolling friction, so you can ignore air drag in this problem.
   a. What propulsion force is needed to keep the car moving forward at a constant speed?
   b. What power is required for propulsion at 5.0 m/s?
   c. If the car gets 15 mpg when driving on sand, what is the car's efficiency?

66. Air conditioners sold in the United States are given a seasonal energy-efficiency ratio (SEER) rating that consumers can use to compare different models. A SEER rating is the ratio of heat pumped to energy input, similar to a COP but using English units, so a higher SEER rating means a more efficient model. You can determine the COP of an air conditioner by dividing the SEER rating by 3.4. For typical inside and outside temperatures when you'd be using air conditioning, estimate the theoretical maximum SEER rating of an air conditioner. (New air conditioners must have a SEER rating that exceeds 13, quite a bit less than the theoretical maximum, but there are practical issues that reduce efficiency.)

67. The surface waters of tropical oceans are at a temperature of 27°C while water at a depth of 1200 m is at 3°C. It has been suggested these warm and cold waters could be the energy reservoirs for a heat engine, allowing us to do work or generate electricity from the thermal energy of the ocean. What is the maximum efficiency possible of such a heat engine?

68. ‖ The light energy that falls on a square meter of ground over the course of a typical sunny day is about 20 MJ. The average rate of electric energy consumption in one house is 1.0 kW.
   a. On average, how much energy does one house use during each 24 h day?
   b. If light energy to electric energy conversion using solar cells is 15% efficient, how many square miles of land must be covered with solar cells to supply the electrical energy for 250,000 houses? Assume there is no cloud cover.

Passage Problems

Kangaroo Locomotion BIO

Kangaroos have very stout tendons in their legs that can be used to store energy. When a kangaroo lands on its feet, the tendons stretch, transforming kinetic energy of motion to elastic potential energy. Much of this energy can be transformed back into kinetic energy as the kangaroo takes another hop. The kangaroo's peculiar hopping gait is not very efficient at low speeds but is quite efficient at high speeds.

Figure P11.69 shows the energy cost of human and kangaroo locomotion. The graph shows oxygen uptake (in mL/s) per kg of body mass, allowing a direct comparison between the two species.

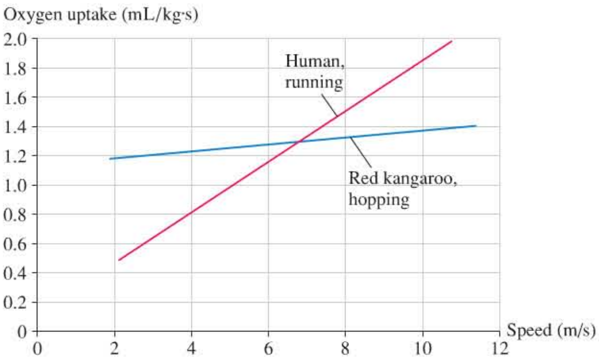


FIGURE P11.69 Oxygen uptake (a measure of energy use per second) for a running human and a hopping kangaroo.

For humans, the energy used per second (i.e., power) is proportional to the speed. That is, the human curve nearly passes through the origin, so running twice as fast takes approximately twice as much power. For a hopping kangaroo, the graph of energy use has only a very small slope. In other words, the energy used per second changes very little with speed. Going faster requires very little additional power. Treadmill tests on kangaroos and observations in the wild have shown that they do not become winded at any speed at which they are able to hop. No matter how fast they hop, the necessary power is approximately the same.

69. | A person runs 1 km. How does his speed affect the total energy needed to cover this distance?
   A. A faster speed requires less total energy.
   B. A faster speed requires more total energy.
   C. The total energy is about the same for a fast speed and a slow speed.

70. | A kangaroo hops 1 km. How does its speed affect the total energy needed to cover this distance?
   A. A faster speed requires less total energy.
   B. A faster speed requires more total energy.
   C. The total energy is about the same for a fast speed and a slow speed.

71. | At a speed of 4 m/s,
   A. A running human is more efficient than an equal-mass hopping kangaroo.
   B. A running human is less efficient than an equal-mass hopping kangaroo.
   C. A running human and an equal-mass hopping kangaroo have about the same efficiency.

72. | At approximately what speed would a human use half the power of an equal-mass kangaroo moving at the same speed?
   A. 3 m/s B. 4 m/s C. 5 m/s D. 6 m/s

73. | At what speed does the hopping motion of the kangaroo become more efficient than the running gait of a human?
   A. 3 m/s B. 5 m/s C. 7 m/s D. 9 m/s

STOP TO THINK ANSWERS

Stop to Think 11.1: C. In each case, what you get is the potential-energy change of the box. Crane 2 lifts a box with twice the mass the same distance as crane 1, so you get twice as much energy with crane 2. How about what you have to pay? Crane 2 uses 20 kJ, crane 1 only 10 kJ. Comparing crane 1 and crane 2, we find crane 2 has twice the energy out for twice the energy in, so the efficiencies are the same.

Stop to Think 11.2: C. As the body uses chemical energy from food, approximately 75% is transformed into thermal energy. Also, kinetic energy of motion of the legs and feet is transformed into thermal energy with each stride. Most of the chemical energy is transformed into thermal energy.

Stop to Think 11.3: C. Samples 1 and 2 have the same thermal energy, which is the total kinetic energy of all the atoms. Sample 1 has twice as many atoms, so the average energy per atom, and thus the temperature, must be less.

Stop to Think 11.4: C. The radiator is at a higher temperature than the surrounding air. Thermal energy is transferred out of the system to the environment, so $Q < 0$.

Stop to Think 11.5: A, D. The efficiency is fixed by the ratio of T_C to T_H. Decreasing this ratio increases efficiency; the heat engine will be more efficient with a hotter hot reservoir or a colder cold reservoir.

Stop to Think 11.6: A, D. The closer the temperatures of the hot and cold reservoirs, the more efficient the heat pump can be. (It is also true that having the two temperatures be closer will cause less thermal energy to "leak" out.) Any change that makes the two temperatures closer will allow the refrigerator to use less energy to run.

Stop to Think 11.7: B. In this case, kinetic energy is transformed into potential energy; there is no entropy change. In the other cases, energy is transformed into thermal energy, meaning entropy increases.

PART II SUMMARY

Conservation Laws

In Part II we have discovered that we don't need to know all the details of an interaction to relate the properties of a system "before" the interaction to those "after" the interaction. We also found two important quantities, momentum and energy, that are often conserved. Momentum and energy are characteristics of a system.

Momentum and energy have conditions under which they are conserved. The total momentum $\vec{P}$ and the total energy E are conserved for an *isolated system*. Of course, not all systems are isolated. For both momentum and energy, it was useful to develop a *model* of a system interacting with its environment. Interactions within the system do not change $\vec{P}$ or E. The kinetic, potential, and thermal energies *within* the system can be transformed without changing E. Interactions between the system and the environment *do* change the system's momentum and energy. In particular:

- Impulse is the transfer of momentum to or from the system: $\Delta\vec{p} = \vec{J}$.
- Work is the transfer of energy to or from the system in a mechanical interaction: $\Delta E = W$.
- Heat is the energy transferred to or from the system in a thermal interaction: $\Delta E = Q$.

The laws of conservation of momentum and energy, when coupled with the laws of Newtonian mechanics of Part I, form a powerful set of tools for analyzing motion. But energy is a concept that can be used for much more than the study of motion; it can be used to analyze how your body uses food, how the sun shines, and a host of other problems.

The study of the transformation of energy from one form into another reveals certain limits. Thermal energy is different from other forms of energy. A transformation of energy from another form into thermal energy is *irreversible*. The study of heat and thermal energy thus led us to the discipline of thermodynamics, a subject we will take up further in Part III as we look at properties of matter.

KNOWLEDGE STRUCTURE II Conservation Laws

BASIC GOALS

How is the system "after" an interaction related to the system "before"?
What quantities are conserved, and under what conditions?
Why are some energy changes more efficient than others?

GENERAL PRINCIPLES

Law of conservation of momentum For an isolated system, $\vec{P}_f = \vec{P}_i$.
Law of conservation of energy For an isolated system, there is no change in the system's energy:

$$\Delta K + \Delta U_g + \Delta U_s + \Delta E_{th} + \Delta E_{chem} + \cdots = 0$$

Energy can be exchanged with the environment as work or heat:

$$\Delta K + \Delta U_g + \Delta U_s + \Delta E_{th} + \Delta E_{chem} + \cdots = W + Q$$

Laws of thermodynamics First law: If only thermal energy changes, $\Delta E_{th} = W + Q$.
Second law: The entropy of an isolated system always increases.

BASIC PROBLEM-SOLVING STRATEGY

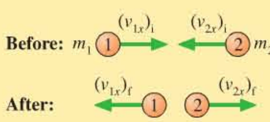
Draw a visual overview for the system "before" and "after"; then use the conservation of momentum or energy equations to relate the two. If necessary, calculate impulse and/or work.

Momentum and impulse

In a collision, the total momentum

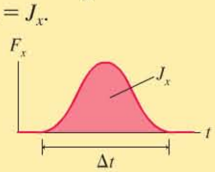
$$\vec{P} = \vec{p}_1 + \vec{p}_2 = m_1\vec{v}_1 + m_2\vec{v}_2$$

is the same before and after.

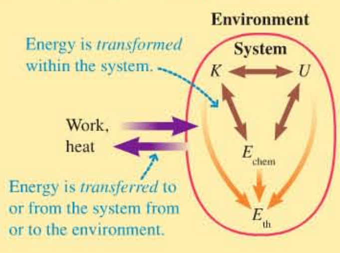


A force can change the momentum of an object. The change is the **impulse:** $\Delta p_x = J_x$.

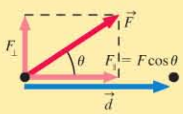
$J_x =$ area under force curve



Basic model of energy

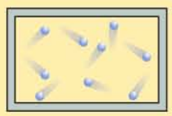


Work $W = F_\parallel d$ is done by the component of a force parallel to a displacement.



Limitations on energy transfers and transformations

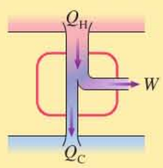
Thermal energy is random kinetic energy. Changing other forms of energy to thermal energy is **irreversible.**



When transforming energy from one form into another, some may be "lost" as thermal energy. This limits efficiency:

$$\text{efficiency: } e = \frac{\text{what you get}}{\text{what you had to pay}}$$

A heat engine can convert thermal energy to useful work. The efficiency must be less than 100%.



Order Out of Chaos

The second law of thermodynamics specifies that "the future" is the direction of entropy increase. But, as we have seen, this doesn't mean that systems must invariably become more random. You don't need to look far to find examples of systems that spontaneously evolve to a state of greater order.

A snowflake is a perfect example. As water freezes, the random motion of water molecules is transformed into the orderly arrangement of a crystal. The entropy of the snowflake is less than that of the water vapor from which it formed. Has the second law of thermodynamics been turned on its head?

The entropy of the water molecules in the snowflake certainly decreases, but the water doesn't freeze as an isolated system. For it to freeze, heat energy must be transferred from the water to the surrounding air. The entropy of the air increases by *more* than the entropy of the water decreases. Thus the *total* entropy of the water + air system increases when a snowflake is formed, just as the second law predicts. If the system isn't isolated, its entropy can decrease without violating the second law as long as the entropy increases somewhere else.

Systems that become *more* ordered as time passes, and in which the entropy decreases, are called *self-organizing systems.* These systems can't be isolated. It is common in self-organizing systems to find a substantial flow of energy *through* the system. Your body takes in chemical energy from food, makes use of that energy, and then gives waste heat back to the environment. It is this energy flow that allows systems to develop a high degree of order and a very low entropy. The entropy of the environment undergoes a significant *increase* so as to let selected subsystems decrease their entropy and become more ordered.

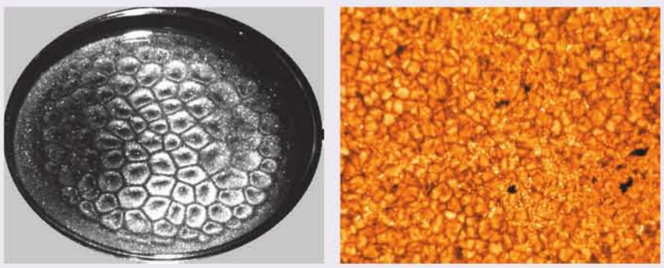
Self-organizing systems don't violate the second law of thermodynamics, but this fact doesn't really explain their existence. If you toss a coin, no law of physics says that you can't get heads 100 times in a row—but you don't expect this to happen. Can we show that self-organization isn't just possible, but likely?

Let's look at a simple example. Suppose you heat a shallow dish of oil at the bottom, while holding the temperature of the top constant. When the temperature difference between the top and the bottom of the dish is small, heat is transferred from the bottom to the top by conduction. But convection begins when the temperature difference becomes large enough. The pattern of convection needn't be random, though; it can develop in a stable, highly ordered pattern, as we see in the figure. Convection is a much more efficient means of transferring energy than conduction, so the rate of transfer is *increased* as a result of the development of these ordered *convection cells.*

The development of the convection cells is an example of self-organization. The roughly 10^{23} molecules in the fluid had been moving randomly but now have begun behaving in a very orderly fashion. But there is more to the story. The convection cells transfer energy from the hot lower side of the dish to the cold upper side. This hot-to-cold energy transfer increases the entropy of the surrounding environment, as we have seen. In becoming more organized, the system has become more effective at transferring heat, resulting in a greater rate of entropy increase! Order has arisen out of disorder in the system, but the net result is a more rapid increase of the disorder of the universe.

Convection cells are thus a thermodynamically favorable form of order. We should expect this, because convection cells aren't confined to the laboratory. We see them in the sun, where they transfer energy from lower levels to the surface, and in the atmosphere of the earth, where they give rise to some of our most dramatic weather.

Self-organizing systems are a very active field of research in physical and biological sciences. The 1977 Nobel Prize in chemistry was awarded to the Belgian scientist Ilya Prigogine for his studies of *nonequilibrium thermodynamics,* the basic science underlying self-organizing systems. Prigogine and others have shown how energy flow through a system can, when the conditions are right, "bring order out of chaos." And this spontaneous ordering is not just possible—it can be probable. The existence and evolution of self-organizing systems, from thunderstorms to life on earth, might just be nature's preferred way of increasing entropy in the universe.



Convection cells in a shallow dish of oil heated from below (left) and in the sun (right). In both, warmer fluid is rising (lighter color) and cooler fluid is sinking (darker color).

The following questions are related to the passage "Order Out of Chaos" on the previous page.

1. When water freezes to make a snowflake crystal, the entropy of the water
   A. Decreases.
   B. Increases.
   C. Does not change.
2. When thermal energy is transferred from a hot object to a cold object, the overall entropy
   A. Decreases.
   B. Increases.
   C. Does not change.
3. Do convection cells represent a reversible process?
   A. Yes, because they are orderly.
   B. No, because they transfer thermal energy from hot to cold.
   C. It depends on the type of convection cell.
4. In an isolated system far from thermal equilibrium, as time passes,
   A. The total energy stays the same; the total entropy stays the same.
   B. The total energy decreases; the total entropy increases.
   C. The total energy stays the same; the total entropy increases.
   D. The total energy decreases; the total entropy stays the same.

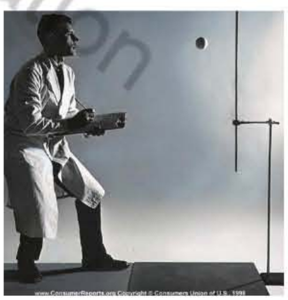
The following passages and associated questions are based on the material of Part II.

Big Air

A new generation of pogo sticks lets a rider bounce more than 2 meters off the ground by using elastic bands to store energy. When the pogo's plunger hits the ground, the elastic bands stretch as the pogo and rider come to rest. At the low point of the bounce, the stretched bands start to contract, pushing out the plunger and launching the rider into the air. For a total mass of 80 kg (rider plus pogo), a stretch of 0.40 m launches a rider 2.0 m above the starting point.

5. If you were to jump to the ground from a height of 2 meters, you'd likely injure yourself. But a pogo rider can do this repeatedly, bounce after bounce. How does the pogo stick make this possible?
   A. The elastic bands absorb the energy of the bounce, keeping it from hurting the rider.
   B. The elastic bands warm up as the rider bounces, absorbing dangerous thermal energy.
   C. The elastic bands simply convert the rider's kinetic energy to potential energy.
   D. The elastic bands let the rider come to rest over a longer time, meaning less force.
6. Assuming that the elastic bands stretch and store energy like a spring, how high would the 80 kg pogo and rider go for a stretch of 0.20 m?
   A. 2.0 m B. 1.5 m C. 1.0 m D. 0.50 m
7. Suppose a much smaller rider (total mass of rider plus pogo of 40 kg) mechanically stretched the elastic bands of the pogo by 0.40 m, then got on the pogo and released the bands. How high would this unwise rider go?
   A. 8.0 m B. 6.0 m C. 4.0 m D. 3.0 m
8. A pogo and rider of 80 kg total mass at the high point of a 2.0 m jump will drop 1.6 m before the pogo plunger touches the ground, slowing to a stop over an additional 0.40 m as the elastic bands stretch. What approximate average force does the pogo stick exert on the ground during the landing?
   A. 4000 N B. 3200 N C. 1600 N D. 800 N
9. Riders can use fewer elastic bands, reducing the effective spring constant of the pogo. The maximum stretch of the bands is still 0.40 m. Reducing the number of bands will
   A. Reduce the force on the rider and give a lower jump height.
   B. Not change the force on the rider but give a lower jump height.
   C. Reduce the force on the rider but give the same jump height.
   D. Make no difference to the force on the rider or the jump height.

Testing Tennis Balls



Tennis balls are tested by being dropped from a height of 2.5 m onto a concrete floor. The 57 g ball hits the ground, compresses, then rebounds. A ball will be accepted for play if it rebounds to a height of about 1.4 m; it will be rejected if the bounce height is much more or much less than this.

10. Consider the sequence of energy transformations in the bounce. When the dropped ball is motionless on the floor, compressed, and ready to rebound, most of the energy is in the form of
    A. Kinetic energy.
    B. Gravitational potential energy.
    C. Thermal energy.
    D. Elastic potential energy.
11. If a ball is "soft," it will spend more time in contact with the floor and won't rebound as high as it is supposed to. The force on the floor of the "soft" ball is ________ the force on the floor of a "normal" ball.
    A. Greater than
    B. The same as
    C. Less than
12. Suppose a ball is dropped from 2.5 m and rebounds to 1.4 m.
    a. How fast is the ball moving just before it hits the floor?
    b. What is the ball's speed just after leaving the floor?
    c. What happens to the "lost" energy?
    d. If the time of the collision with the floor is 6.0 ms, what is the average force on the ball during the impact?

Squid Propulsion BIO

Squid usually move by using their fins, but they can utilize a form of "jet propulsion," ejecting water at high speed to rocket them backward, as shown in Figure II.1. A 4.0 kg squid can slowly draw in and then quickly eject 0.30 kg of water. The water is ejected in 0.10 s at a speed of 10 m/s. This gives the squid a quick burst of speed to evade predators or catch prey.

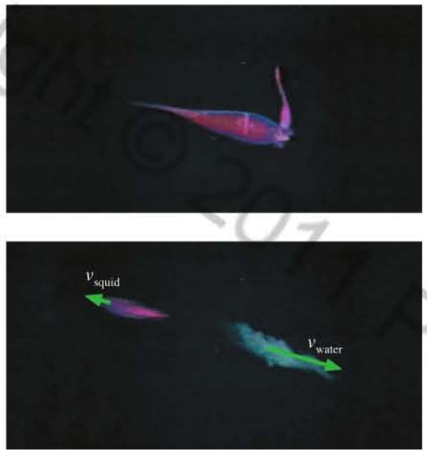


FIGURE II.1

13. What is the speed of the squid immediately after the water is ejected?
 A. 10 m/s
 B. 7.5 m/s
 C. 1.3 m/s
 D. 0.75 m/s
14. What is the squid's approximate acceleration in g?
 A. $10g$ B. $7.5g$ C. $1.0g$ D. $0.75g$
15. What is the average force on the water during the jet?
 A. 100 N B. 30 N C. 10 N D. 3.0 N
16. This form of locomotion is speedy, but is it efficient? The energy that the squid expends goes two places: the kinetic energy of the squid and the kinetic energy of the water. Think about how to define "what you get" and "what you had to pay"; then calculate an efficiency for this particular form of locomotion. (You can ignore biomechanical efficiency for this problem.)

Teeing Off

A golf club has a lightweight flexible shaft with a heavy block of wood or metal (called the head of the club) at the end. A golfer making a long shot off the tee uses a driver, a club whose 300 g head is much more massive than the 46 g ball it will hit. The golfer swings the driver so that the club head is moving at 40 m/s just before it collides with the ball. The collision is so rapid that it can be treated as the collision of a moving 300 g mass (the club head) with a stationary 46 g mass (the ball); the shaft of the club and the golfer can be ignored. The collision takes 5.0 ms, and the ball leaves the tee with a speed of 69 m/s.

17. What is the change in momentum of the ball during the collision?
 A. 1.4 kg·m/s B. 1.8 kg·m/s
 C. 3.2 kg·m/s D. 5.1 kg·m/s
18. What is the speed of the club head immediately after the collision?
 A. 29 m/s B. 25 m/s C. 19 m/s D. 11 m/s
19. Is this a perfectly elastic collision?
 A. Yes
 B. No
 C. There is insufficient information to make this determination.
20. If we define the kinetic energy of the club head before the collision as "what you paid" and the kinetic energy of the ball immediately after as "what you get," what is the efficiency of this energy transfer?
 A. 0.54 B. 0.46 C. 0.37 D. 0.27

Additional Integrated Problems

21. Football players measure their acceleration by seeing how fast they can sprint 40 yards (37 m). A zippy player can, from a standing start, run 40 yards in 4.1 s, reaching a top speed of about 11 m/s. For an 80 kg player, what is the average power output for this sprint?
 A. 300 W B. 600 W C. 900 W D. 1200 W
22. The unit of horsepower was defined by considering the power output of a typical horse. Working-horse guidelines in the 1900s called for them to pull with a force equal to 10% of their body weight at a speed of 3.0 mph. For a typical working horse of 1200 lb, what power does this represent in W and in hp?
23. A 100 kg football player is moving at 6.0 m/s to the east; a 130 kg player is moving at 5.0 m/s to the west. They meet, each jumping into the air and grabbing the other player. While they are still in the air, which way is the pair moving, and how fast?
24. A swift blow with the hand can break a pine board. As the hand hits the board, the kinetic energy of the hand is transformed into elastic potential energy of the bending board; if the board bends far enough, it breaks. Applying a force to the center of a particular pine board deflects the center of the board by a distance that increases in proportion to the force. Ultimately the board breaks at an applied force of 800 N and a deflection of 1.2 cm.
 a. To break the board with a blow from the hand, how fast must the hand be moving? Use 0.50 kg for the mass of the hand.
 b. If the hand is moving this fast and comes to rest in a distance of 1.2 cm, what is the average force on the hand?
25. A child's sled has rails that slide with little friction across the snow. Logan has an old wooden sled with heavy iron rails that has a mass of 10 kg—quite a bit for a 30 kg child! Logan runs at 4.0 m/s and leaps onto the stationary sled and holds on tight as it slides forward. The impact time with the sled is 0.25 s.
 a. Immediately after Logan jumps on the sled, how fast is it moving?
 b. What was the force on the sled during the impact?
 c. How much energy was "lost" in the impact? Where did this energy go?