Although the heat pump in this example is far less expensive, two factors make the comparison unrealistic. First, real heat pumps operate at a COP closer to half their theoretical maximum, approximately doubling the cost over that of this example. Second, the initial installation of a heat pump is very expensive, so that economic considerations may still favor oil. As heat pump technology advances, though, these devices will come into wider use.

If we are concerned about energy and not just cost, we should also consider the source of electricity to run the heat pump. If this is a thermal power plant operating at an efficiency of, say, $\frac{1}{3}$, then three times as much energy is used at the power plant as is supplied to the heat pump. The overall efficiency of the heat pump and power plant may actually be less than that of the oil furnace (see Problem 22).

19.4
THE THERMODYNAMIC TEMPERATURE SCALE

When we discussed temperature measurement in Chapter 16, we encountered the problem that no single method works at all temperatures. We settled on the gas thermometer as the best laboratory standard, but even a helium gas thermometer becomes useless at temperatures below about 1 K.

The operation of a reversible heat engine provides a way of measuring temperature that is independent of the particular device used and that works at any temperature. We can take such an engine—or any thermodynamic system—through a complete cycle of reversible processes. Measuring the heats Q_c and Q_h, or one of these and the work output, allows us to calculate the efficiency and from it the temperature ratio. If one temperature is known, we can then determine the other. Although this may seem a rather obscure way to measure temperature, it is actually used in certain experiments at very low temperature. And it provides an absolute way of defining temperature that can, in principle, be used at any temperature.

thermodynamic temperature scale

The temperature scale so defined is called the **thermodynamic temperature scale.**

absolute zero

The zero of the thermodynamic scale—**absolute zero**—is absolute in that it represents a state of maximum order. This is reflected in the fact that a heat engine rejecting heat at absolute zero would operate at 100 per cent efficiency, as suggested by Equation 19–5 with $T_c = 0$. Unfortunately, there are no heat reservoirs at absolute zero, and it can be shown that it is impossible to cool anything to this temperature in a finite number of steps. The statement that it is impossible to reach absolute zero in a finite number

third law of thermodynamics

of steps is called the **third law of thermodynamics.**

19.5
ENTROPY AND THE QUALITY OF ENERGY

If offered a joule of energy, would you rather have that joule delivered in the form of mechanical work, heat transferred from an object at 1000 K, or heat transferred from an object at 300 K? Your answer might depend on what you wanted to do with the energy. If you wanted to lift or accelerate a mass, then you would be smart to take your energy as work. But if you

want to keep warm, then heat from the 300 K object would be perfectly acceptable.

If, on the other hand, you weren't sure what you wanted to do with the energy, then which should you take? The second law of thermodynamics makes the answer clear: you should take the work. Why? Because you could use it directly to give one joule of mechanical energy—kinetic or potential—to a mass, or you could, through friction or other irreversible processes, use it to raise the temperature of something.

If you chose heat from the 300 K object as the form for your joule of energy, then you could supply a full joule only to objects cooler than 300 K. You could not do mechanical work unless you used your joule of 300 K heat to run a heat engine. With its maximum temperature only a little above ambient, your engine would be very inefficient, and you could only extract a small fraction of a joule of mechanical energy. Nor could you supply your energy to anything hotter than 300 K, unless you ran a refrigerator—and that would take additional work over and above your one joule. You would be better off with one joule of energy in the form of 1000 K heat, for you could transfer it to anything cooler than 1000 K, or could run a heat engine to produce a substantial fraction of a joule of mechanical energy.

Taking your one joule of energy in the form of work gives you the most options. Anything you can do with a joule of energy, you can do with the work. Energy taken as heat is less versatile, with the joule of 300 K energy the least useful of the three. We are not talking here about the quantity of energy—we have exactly one joule in each case—but about the **quality of energy,** indicated by the ability to do a variety of useful tasks. Schematically, we can describe energy quality on a diagram like Fig. 19–22. In Fig. 19–22, we list electricity with mechanical energy because electrical energy flows without requiring a temperature difference, and thus qualifies as work. We can readily convert an entire amount of energy from higher to lower quality, but the second law of thermodynamics prevents us from going in the opposite direction with 100 per cent efficiency.

The notion of energy quality has important implications for our efficient use of energy, suggesting that we try to match the quality of available energy sources to our energy needs. Figure 19–23 shows a breakdown of United States energy use by quality; it makes clear that much of our energy demand is for relatively low-quality heat. Although we can do anything we want with a high-quality energy source like flowing water or electricity, it makes sense to put our high-quality energy to high-quality uses, and use lower-quality sources to meet other needs. For example, we often heat water with electricity—a great convenience to the homeowner, but a thermodynamic folly, for in so doing we convert the highest-quality form of energy into low-grade heat. If our electricity comes from a thermal power plant, we already threw away two-thirds of the energy from fuel as waste heat. It makes little sense to run an elaborate and inefficient heat engine only to have its high-quality output of electricity used for low-temperature heating. On the other hand, if we heat our water by burning a fuel directly at the water heater, then we can, in principle, transfer all the fuel's energy to the water. Another approach to matching energy quality with energy needs is to combine the production of steam for heating with the generation of elec-
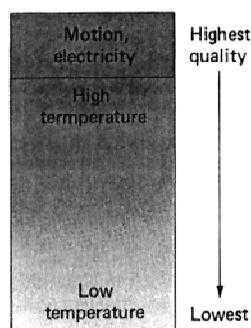

Fig. 19–22
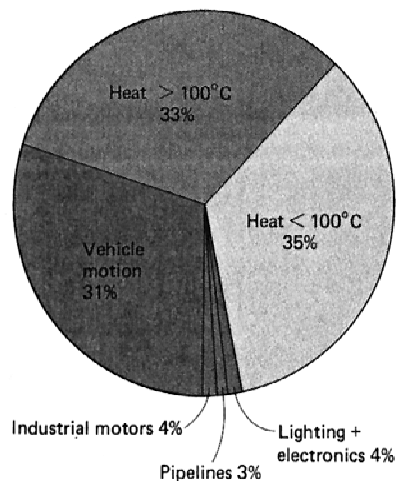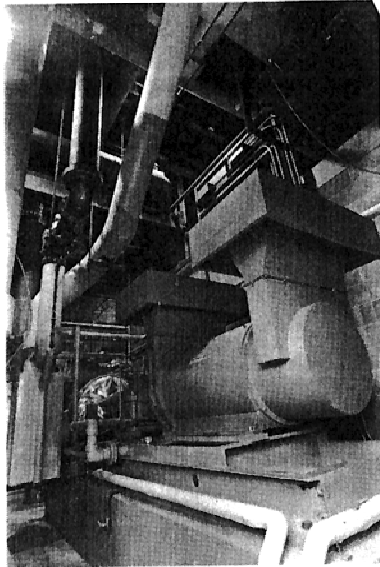Energy quality.

quality of energy

Fig. 19–23
Energy use in the United States, organized by energy quality.

cogeneration

tricity. Popular in Europe for many years, this idea of **cogeneration** is gaining popularity in the United States (Fig. 19–24).

Entropy

How can we make more precise the notion of energy quality? Consider the system shown in Fig. 19–25. It consists of two identical objects, initially at different temperatures and thermally insulated from each other and from the environment. The two objects are then brought into thermal contact, so that they eventually reach equilibrium at the same temperature.

What has changed between the initial and final states of our system? Not the energy, for we did no work on the system, and lost no heat, so that the total energy remained constant. But we have lost the ability to do useful work. In the initial state, we could have run a heat engine between the hot and cool objects. In the final state, there is no temperature difference and therefore no way to run a heat engine. The quality of the energy has decreased. Can we find some system property that reflects this change in energy quality?
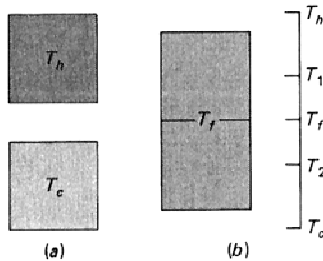
Figure 19–26 shows another example of a system whose energy quality deteriorates without any loss of total energy. We start with a gas confined by a removable partition to one side of an insulated container; the other side is evacuated. We then remove the partition and allow the gas to expand freely to fill the whole container. Since the process is adiabatic, and no work is done, the total energy does not change. But we lose the ability to extract some of that energy as useful work. Starting with the initial state, we could have opened a small hole and used the escaping gas to run a turbine (Fig. 19–27). In the final state, there is no way to run the turbine, for there is no pressure difference across which the gas will flow. Again, can we find a quantity that describes this decrease in energy quality?

To find such a quantity, we consider an ideal gas undergoing a Carnot cycle (Fig. 19–28). In deriving the efficiency of such a cycle, we found that

$$\frac{Q_c}{Q_h} = \frac{T_c}{T_h}, \tag{19–8}$$

where Q_c is defined as the heat *rejected from* the system to the low-temperature reservoir at T_c, and Q_h is the heat *added to* the system from the high-temperature reservoir at T_h.

We now change our convention on the meaning of Q_c, taking it also to mean heat *added to* our system. Changing this convention changes the sign of Q_c, for positive heat rejected from our system is the same as negative heat added to it. With this change in convention, Equation 19–8 may be written

Fig. 19–24
This cogenerating unit at Middlebury College provides 600 kW of electricity from high-pressure steam produced in the college's heating plant. After passing through a turbine that drives the electric generator, the steam is used to heat buildings.
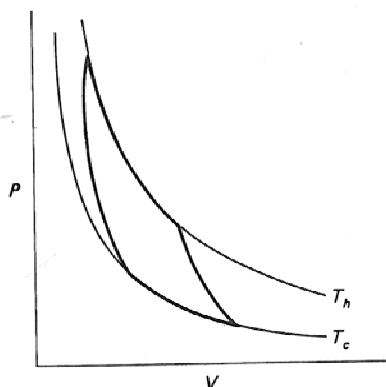

Fig. 19–25
Two identical objects, initially at different temperatures, are allowed to come to thermal equilibrium.

Fig. 19–26
Adiabatic free expansion. No energy leaves the system during the process, but energy quality deteriorates.
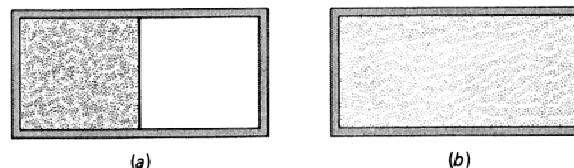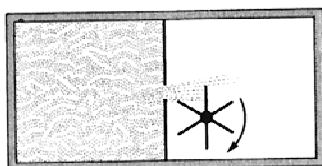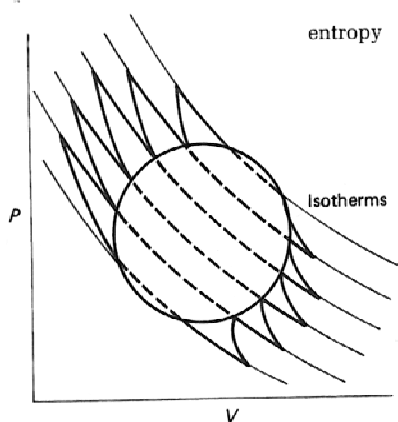
Fig. 19-27
A means of extracting useful work from the gas in the initial state of Fig. 19-26.

Fig. 19-28
Carnot cycle for an ideal gas.

$$\frac{Q_c}{T_c} + \frac{Q_h}{T_h} = 0,$$

or

$$\sum \frac{Q}{T} = 0. \tag{19-9}$$

Although this result follows specifically from consideration of a Carnot cycle consisting of two adiabatic and two isothermal steps, we can generalize it to any reversible cycle. Any such cycle can be approximated arbitrarily closely by a sequence of adiabatic and isothermal steps (Fig. 19-29). One way to move around our reversible cycle is to take in the entire sequence of Carnot cycles shown in Fig. 19-29. But Equation 19-9 holds for each of these Carnot cycles, so that we must have

$$\sum \frac{Q}{T} = 0 \tag{19-10}$$

over the entire set of Carnot cycles. Although the Q's in this expression are heats added in the isothermal sections of the individual Carnot cycles, those parts of a Carnot cycle that traverse the interior of our closed path are crossed twice, in opposite directions (Fig. 19-30), so that their heats cancel. Then Equation 19-10 applies equally well to the jagged approximation to our closed path. In the limit as we approximate the smooth curve ever closer by more and more adiabatic and isothermal segments, Equation 19-10 becomes

$$\oint \frac{dQ}{T} = 0, \tag{19-11}$$

where the circle on the integral sign indicates that we integrate over all the heat transfers dQ along a *closed* path.

Equation 19-11 tells us that there is some quantity—a small amount of which is given by dQ/T—that does not change when we traverse a closed path, bringing our system back to its original state. We call this quantity the **entropy**, S, and write
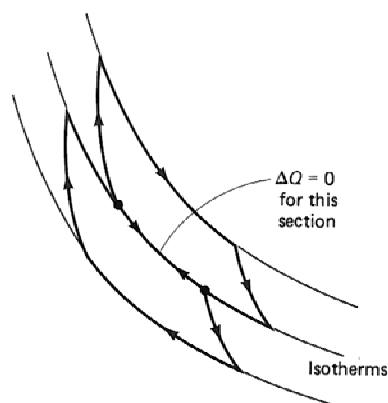
entropy

$$dS = \frac{dQ}{T} \tag{19-12}$$

and

$$S_2 - S_1 = \int_1^2 dS = \int_1^2 \frac{dQ}{T}, \tag{19-13}$$

where the limits on the integral refer to two thermodynamic states and S_1 and S_2 are the entropies of those states. Note that we are not concerned with the actual value of the entropy, but only with the *change* in entropy from one state to another. In this way, entropy is like potential energy, for which only changes are meaningful.[*]

Fig. 19-29
An arbitrary cycle may be approximated by a sequence of adiabatic and isothermal steps.

Isotherms

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[*]Actually, the zero of entropy is meaningful, in that zero entropy for any system corresponds to a temperature of absolute zero.

Fig. 19–30
Heat transfers cancel as two adjacent
Carnot cycles are traversed across the
interior of the original cycle.
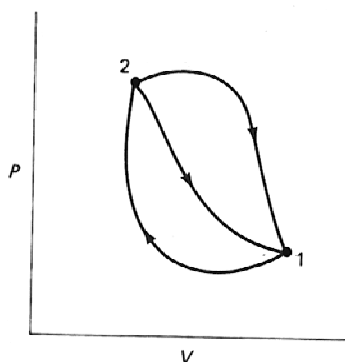
Fig. 19–31
Entropy change is zero when the
system returns to its original state,
so the entropy change from state
2 to state 1 must be independent
of the path taken.

Example 19–5
Entropy

An object of mass m and specific heat c is taken from an initial temperature T_1 to a final temperature T_2. What is the change in its entropy?

Solution

The entropy change is given by Equation 19–13. To evaluate the integral in this equation, we must relate dQ and T. We can do so because we know the specific heat, so that

$$dQ = mcdT.$$

Then Equation 19–13 is

$$\Delta S = \int_{T_1}^{T_2} \frac{dQ}{T} = \int_{T_1}^{T_2} mc \frac{dT}{T} = mc \ln\left(\frac{T_2}{T_1}\right). \qquad (19\text{–}14)$$

What good is this concept of entropy? Equation 19–11 shows that there is no entropy change when we take a system from some initial state and eventually return it to that state. Now suppose we take a system part way around a closed path, from state 1 to state 2 as shown in Fig. 19–31. The entropy change from state 2 back to state 1 must be exactly opposite the change from 1 to 2, so that the change going around the closed path is zero. But this must be true no matter how we get from state 2 to state 1, as suggested by the two possible paths in Fig. 19–31. That is, the entropy change given by Equation 19–13 is independent of path; it is a property only of the initial and final states themselves. Like pressure, temperature, and volume, entropy is a thermodynamic state variable—a quantity that characterizes a given state independently of how a system got into that state.

Equation 19–13 for calculating the entropy difference between two states is meaningful only for reversible processes, for the temperature T is defined only in thermodynamic equilibrium, while an irreversible process takes a system temporarily out of equilibrium. But because entropy is a property of a thermodynamic state, and not of how a system got into that state, entropy changes during an irreversible process can be calculated by finding a reversible process that takes a system between the same two states, then calculating the entropy change for that process using Equation 19–13.

At the beginning of this section, we considered some irreversible processes in which energy quality deteriorated although total energy remained constant. We can come to understand the meaning of entropy by calculating entropy changes for these processes.

Example 19–6
### Adiabatic Free Expansion

An ideal gas, initially confined to one side of an insulated container, is allowed to expand freely to fill the entire container (Fig. 19–32). What is the entropy change for this process?

Solution

Since the free expansion is adiabatic and involves no work, the internal energy and temperature of the ideal gas are unchanged. But the gas volume doubles. To

Fig. 19-32
Adiabatic free expansion.

(a)   (b)
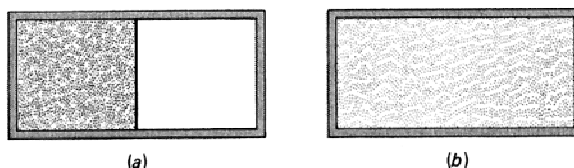
find the entropy change for the irreversible free expansion, we can consider a reversible process taking the system between the same two states. Since the initial and final temperatures are the same, such a process is an isothermal expansion. The heat added during such an expansion is given by Equation 18–6:

$$Q = nRT \ln\left(\frac{V_2}{V_1}\right).$$

Since the temperature is constant during the reversible isothermal expansion, the entropy change of Equation 19–13 is given simply by

$$S_2 - S_1 = \frac{Q}{T} = nR \ln\left(\frac{V_2}{V_1}\right). \tag{19–15}$$

Since the final volume V_2 is larger than the initial volume, the entropy change is positive.

In calculating the entropy change, we used the heat Q that would be required during a reversible isothermal process. Of course, no heat is transferred during the actual adiabatic free expansion—but the entropy change is the same as for the reversible isothermal process, which is why we calculated the entropy change as though such a process had occurred.

Example 19-7
Irreversible Heat Transfer

An object at a high temperature T_h is placed in contact with another object at a lower temperature T_c. What is the change in the entropy of the system when the two have reached equilibrium?

Solution

Again, we have an irreversible process. But we can calculate the entropy change by considering any process that takes the two objects from their initial temperatures to the same final temperature T_f. For example, we cool the hot object by placing it in contact with a heat reservoir initially at T_h and gradually lowering its temperature to T_f. The entropy change is given by

$$\Delta S_h = \int_{T_h}^{T_f} \frac{dQ}{T}.$$

Although we could evaluate this integral, let us for the moment note only that there is some average temperature T_1 that lies between T_h and T_f, for which

$$\Delta S_h = \frac{Q_h}{T_1},$$

where Q_h is the heat absorbed by the hot object. Q_h is, in fact, negative since the hot object loses heat as it cools. Similarly, the entropy change of the cool object is

$$\Delta S_c = \int_{T_c}^{T_f} \frac{dQ}{T} = \frac{Q_c}{T_2},$$

where T_2 is some average temperature that lies between T_c and T_f, and Q_c is the heat absorbed by the cool object.

The total entropy change of the system is

$$\Delta S = \Delta S_c + \Delta S_h = \frac{Q_c}{T_2} + \frac{Q_h}{T_1}.$$

Because the two objects are insulated from the outside world, any heat lost by the hot object in the actual irreversible process is gained by the cool object. Therefore, $Q_h = -Q_c$, and the total entropy change becomes

$$\Delta S = \frac{Q_c}{T_2} - \frac{Q_c}{T_1} = Q_c\left(\frac{1}{T_2} - \frac{1}{T_1}\right).$$

Now whatever the value of T_1, it lies *above* the equilibrium temperature T_f. Similarly, T_2 lies *below* T_f, so that $T_2 < T_1$. Furthermore, Q_c is positive, since the cool object absorbs heat. Therefore, the first term in our expression for ΔS is greater than the second term, so that the entropy change is positive.

Problem 35 explores this example more quantitatively, but the significant conclusion remains: entropy increases during an irreversible heat transfer.

The systems in Examples 19–6 and 19–7 were closed systems, in that they did not exchange energy or matter with their environments. For each system, entropy increased as the system underwent an irreversible process. At the same time, energy that was originally available to do useful work became unavailable. This correlation between an increase in entropy and a decrease in energy quality can be made more quantitative. Consider the free expansion of Example 19–6. Had we let the gas expand isothermally but reversibly, it could have done an amount of work given by Equation 18–5:

$$W = nRT \ln\left(\frac{V_2}{V_1}\right). \tag{19–16}$$

After the irreversible free expansion, the gas could no longer do this work, even though its energy content was unchanged. During the expansion, its entropy increased by an amount given by Equation 19–15:

$$\Delta S = nR \ln\left(\frac{V_2}{V_1}\right).$$

Comparing this entropy change with the energy that became unavailable to do work (Equation 19–16), we have

$$E_{\text{unavailable}} = T\Delta S. \tag{19–17}$$

Equation 19–17 is an example of a more general relation between entropy and the quality or availability of energy:

entropy and the quality of energy

> During an irreversible process in which the entropy of a system increases by ΔS, an amount of energy $E = T_{\min}\Delta S$ becomes unavailable to do work. $T_{\min}$ is the temperature of the coolest reservoir available to the system.

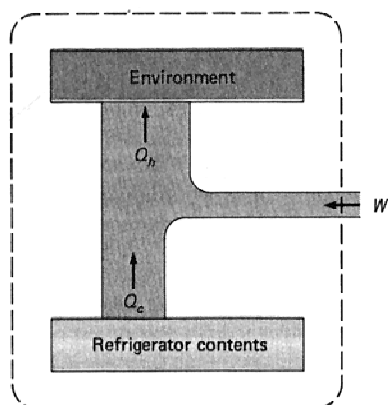This statement shows that entropy provides our desired measure of energy quality. Given two systems that are identical in composition and energy content, the system with the lowest entropy contains the highest-quality energy. An increase in entropy always corresponds to a degradation in energy quality, in that some energy becomes unavailable to do work. Problem 36 explores this energy degradation quantitatively.

Entropy and the Second Law

At the beginning of this chapter, we argued that processes in nature always occur irreversibly, going from ordered states to disordered states. It is this loss of order in thermodynamic systems—as when a hot and a cold object eventually reach the same temperature—that is responsible for energy becoming unavailable to do work. Entropy is a measure of this disorder. Given the tendency of systems to evolve toward disordered states, we can make a statement about entropy that is, in fact, a general statement of the second law of thermodynamics:

second law of thermodynamics

> The entropy of a closed system can never decrease.

At best, the entropy of a closed system remains constant—and this happens only in an ideal, reversible process. If anything irreversible occurs—a slight amount of friction or a deviation from exact thermodynamic equilibrium—then entropy increases. There is no going back. As entropy increases, some energy becomes unavailable to do work, and nothing within the closed system can restore that energy to its original quality. This statement of the second law in terms of entropy is equivalent to our previous statements about the impossibility of perfect heat engines and refrigerators, for the operation of either device would require a decrease in entropy.

What about a system that is not closed? Can't we decrease its entropy? Yes—but only by supplying high-quality energy from outside. When we run a refrigerator, we decrease the entropy of the refrigerator's contents and its surroundings (Fig. 19–33). At the same time, however, we must do work on the refrigerator to effect this entropy decrease—to make heat flow in the direction it doesn't normally go. If we enlarge our closed system to include the power plant or whatever else supplies the refrigerator with work, we would find that the entropy of this new closed system does not decrease. The entropy decrease at the refrigerator is offset by entropy increases elsewhere in the system (Fig. 19–34). If irreversible processes occur anywhere in the system, then there is a net entropy increase.

Fig. 19–33
The entropy decrease of the refrigerator's contents can occur only when work is done on the system from outside.
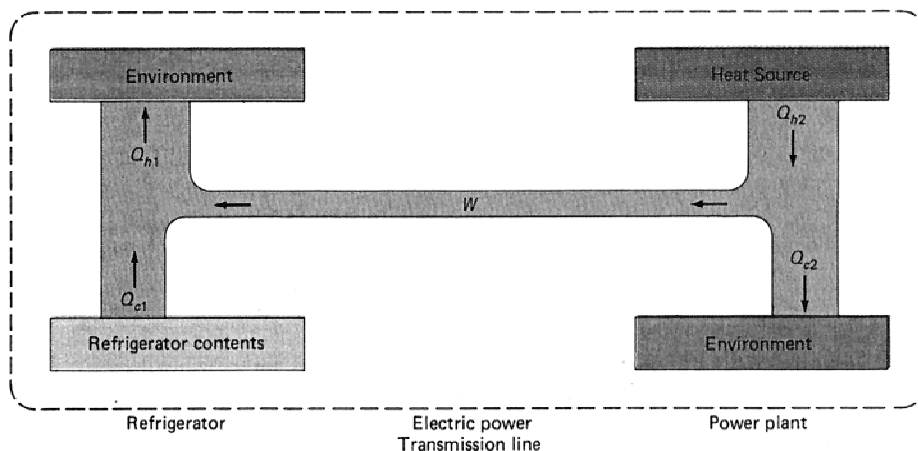
Fig. 19–34
When the source of work for the refrigerator is included in the system, then entropy of the entire system can at best remain constant.

Whenever we encounter a system whose entropy seems to decrease—that gets more rather than less organized—we do not have a closed system. If we enlarge the boundaries of our system to encompass anything that might exchange energy with it, then we will always find that the enlarged system undergoes an entropy increase. Ultimately, we can enlarge our system to include the entire universe. When we do, we have the ultimate statement of the second law of thermodynamics:

second law of thermodynamics

> The entropy of the universe can never decrease.

As last examples of this broad statement, consider the growth of a living thing from the random mix of molecules in its environment, or the construction of a skyscraper from materials that were originally dispersed about the earth, or the appearance of ordered symbols on a printed page from what was a uniformly mixed bottle of ink. All these are processes in which matter goes from near chaos to a highly organized state—akin to separating yolk and white from a scrambled egg. They are certainly processes in which entropy decreases. But earth is not a closed system. It gets high-quality energy from the sun, energy that is ultimately responsible for life and all its actions. If we consider the earth-sun system, then the entropy decrease associated with life and civilization on earth is more than balanced by the entropy increase associated with nuclear fusion inside the sun. We living things represent a remarkable phenomenon—the organization of matter in a universe governed by a tendency toward disorder. But we do not escape the second law of thermodynamics. Our highly organized selves and society, and the entropy decreases they represent, come into being only at the expense of greater entropy increases elsewhere in the universe.

SUMMARY

1. Irreversible processes are those that result in a decrease in organization or order. Friction is an irreversible process that converts directed kinetic energy into the random kinetic energy of molecular motions. Thermodynamic processes that take a nonequilibrium state toward equilibrium are similarly irreversible. Because there are many more ways to arrange matter in a disorganized state than in an organized state, the probability that a disorganized state will become spontaneously organized is impossibly small.

2. The **second law of thermodynamics** comprises a set of equivalent statements to the effect that processes never take matter from less organized to more organized states.

   a. A **heat engine** is a device that extracts heat from a reservoir and converts some of the energy to useful work. Applied to heat engines, the second law states that it is impossible to build a perfect heat engine—one that converts heat to work with one hundred per cent efficiency. The most efficient engine is one that includes only reversible processes. The efficiency of such an engine—of which a **Carnot engine** is one example—is

   $$e = 1 - \frac{T_c}{T_h},$$

   where T_c and T_h are the temperatures of cool and hot reservoirs with which the engine exchanges energy. Because of the second law's limitation on efficiency, practical heat engines like automobile engines and electric power plants exhaust a great deal of waste heat to their environments.

   b. A **refrigerator** is a device that extracts heat from a cool object and transfers it to a warmer object. Applied to refrigerators, the second law states that it is impossible to build a perfect refrigerator—one that transfers heat without energy input from outside. Because of this second-law limitation, real refrigerators require a source of outside power. The efficiency of a refrigerator is measured by its **coefficient**

of **performance,** defined as the ratio of heat extracted from the cool reservoir to the work input to the device:

$$COP = \frac{Q_c}{W}.$$

A refrigerator could be made by running a reversible engine backward, supplying work instead of producing it. Consideration of the second-law efficiency of such an engine then shows that the maximum COP of a refrigerator is

$$COP = \frac{T_c}{T_h - T_c}.$$

3. The second-law limit on the efficiency of a heat engine provides a definition of the **thermodynamic temperature scale.** This scale is independent of any particular type of thermometer or material, and coincides with the ideal gas temperature scale in the range where the latter is useful.
4. **Quality of energy** refers to the ability of a given amount of energy to be converted into work. The highest-quality energy is mechanical or electrical; these are associated with the organized motion or potential energy of particles in an object. Next in quality comes the energy associated with large temperature differences, followed by ever smaller temperature differences. The highest-quality energy can be readily converted to lower quality with one hundred per cent efficiency. The reverse process—conversion of lower-quality to higher-quality energy with one hundred per cent efficiency—is prohibited by the second law of thermodynamics.
5. **Entropy** provides a measure of the relative disorder of a system, so that increasing entropy is associated with decreasing energy quality. Entropy is a thermodynamic state variable, so that its value depends only on the state of a system and not on how the system got into that state. The entropy difference between two states may be calculated by evaluating the expression

$$\Delta S = \int_1^2 \frac{dQ}{T}$$

for any reversible process connecting those states.
6. In terms of entropy, the second law of thermodynamics becomes the statement that the entropy of the universe can never decrease. In this form, the second law is a universal statement about the tendency of systems to evolve toward states of higher disorder.

QUESTIONS

1. Which of the following processes is irreversible?
   a. Stirring sugar into coffee
   b. Building a house
   c. Demolishing a house with a wrecking crane
   d. Demolishing a house by taking it apart piece-by-piece
   e. Warming a bottle of milk by transferring it directly from a refrigerator to stove top
   f. Writing a sentence
   g. Harnessing the energy of falling water to drive machinery
2. Could you cool the kitchen by leaving the refrigerator door open? Explain.
3. Could you heat the kitchen by leaving the oven open? Explain.
4. Why don't we simply refrigerate the cooling water of a power plant before it goes to the plant, thereby increasing the plant's efficiency?
5. Should a car get better mileage in the summer or the winter? Explain.
6. Is there a limit to the maximum temperature that can be achieved by focusing sunlight with a lens? If so, what is it?
7. Name some irreversible processes that occur in a real engine.
8. A power company claims that electric heat is one hundred per cent efficient. Discuss this claim.
9. Steam leaves the turbine of a power plant at 150°C, and is condensed to water at the same temperature by contact with a river at 20°C. What temperature should be used as T_c in evaluating the thermodynamic efficiency?
10. A hydroelectric power plant, using the energy of falling water, can operate with an efficiency arbitrarily close to one hundred per cent. Why?
11. Viewed in isolation, a windmill is not a heat engine, for it converts mechanical energy of wind directly into electrical energy. But taking a broader view, the windmill is part of a natural heat engine. Explain.
12. To maximize the COP of a refrigerator, should you strive for a large or a small temperature difference? Explain.
13. The manufacturer of a heat pump claims that the device will heat your home using only energy already available in the ground. Is this true?
14. Proponents of ocean thermal energy conversion do not seem bothered by the extremely low efficiency of this process. Yet comparably low efficiencies would be intolerable in fossil-fueled or nuclear power plants. Why the difference?
15. Could anything useful be done with the waste heat from a power plant? Give two examples.
16. Does sunlight represent high- or low-quality energy? Explain. (See Question 6.)
17. The heat Q added during the process of adiabatic free expansion is zero. Why can't we then argue from