EXAMPLE 15–7 **ΔU for boiling water to steam.** Determine the change in internal energy of 1.00 liter of water (mass 1.00 kg) at 100°C when it is fully boiled from liquid to gas, which results in 1671 liters of steam at 100°C. Assume the process is done at atmospheric pressure.

APPROACH Our system is the water. The heat required here does not result in a temperature change; rather, a change in phase occurs. We can determine the heat Q required using the latent heat of water, as in Section 14–5. Work too will be done: $W = P\,\Delta V$. The first law of thermodynamics will then give us ΔU.

SOLUTION The latent heat of vaporization of water (Table 14–3) is $L_{\mathrm{V}} = 22.6 \times 10^5\,\mathrm{J/kg}$. So the heat input required for this process is

$$Q = mL = (1.00\,\mathrm{kg})(22.6 \times 10^5\,\mathrm{J/kg}) = 22.6 \times 10^5\,\mathrm{J}.$$

The work done by the water is (Eq. 15–3)

$$W = P\,\Delta V = (1.01 \times 10^5\,\mathrm{N/m^2})\left[(1671 \times 10^{-3}\,\mathrm{m^3}) - (1 \times 10^{-3}\,\mathrm{m^3})\right] = 1.69 \times 10^5\,\mathrm{J},$$

where we used $1\,\mathrm{atm} = 1.01 \times 10^5\,\mathrm{N/m^2}$ and $1\,\mathrm{L} = 10^3\,\mathrm{cm^3} = 10^{-3}\,\mathrm{m^3}$. Then

$$\Delta U = Q - W = (22.6 \times 10^5\,\mathrm{J}) - (1.7 \times 10^5\,\mathrm{J}) = 20.9 \times 10^5\,\mathrm{J}.$$

NOTE Most of the heat added goes to increasing the internal energy of the water (increasing molecular energy to overcome the attraction that held the molecules close together in the liquid state). Only a small part ($< 10\%$) goes into doing work.

EXERCISE D Equation 14–1, $U = \frac{3}{2}nRT$, tells us that $\Delta U = 0$ in Example 15–7 because $\Delta T = 0$. Yet we determined that $\Delta U = 21 \times 10^5\,\mathrm{J}$. What is wrong?

* 15–3 Human Metabolism and the First Law

PHYSICS APPLIED
Energy in the human body

Human beings and other animals do work. Work is done when a person walks or runs, or lifts a heavy object. Work requires energy. Energy is also needed for growth—to make new cells, and to replace old cells that have died. A great many energy-transforming processes occur within an organism, and they are referred to as *metabolism*.

We can apply the first law of thermodynamics,

$$\Delta U = Q - W,$$

to an organism: say, the human body. Work W is done by the body in its various activities; if this is not to result in a decrease in the body's internal energy (and temperature), energy must somehow be added to compensate. The body's internal energy is not maintained by a flow of heat Q into the body, however. Normally, the body is at a higher temperature than its surroundings, so heat usually flows *out* of the body. Even on a very hot day when heat is absorbed, the body has no way of utilizing this heat to support its vital processes. What then is the source of energy that allows us to do work? It is the internal energy (chemical potential energy) stored in foods (Fig. 15–9). In a closed system, the internal energy changes only as a result of heat flow or work done. In an open system, such as a human, internal energy itself can flow into or out of the system. When we eat food, we are bringing internal energy into our bodies directly, which thus increases the total internal energy U in our bodies. This energy eventually goes into work and heat flow from the body according to the first law.

FIGURE 15–9 Bike rider getting an input of energy.