15–1 The First Law of Thermodynamics

In Section 14–2, we defined the internal energy of a system as the sum total of all the energy of the molecules of the system. We would expect that the internal energy of a system would be increased if work was done on the system, or if heat were added to it. Similarly the internal energy would be decreased if heat flowed out of the system or if work were done by the system on something in the surroundings.

Thus it is reasonable to extend the work-energy principle and propose an important law: the change in internal energy of a closed system, ΔU, will be equal to the energy added to the system by heating minus the work done by the system on the surroundings. In equation form we write

$$\Delta U = Q - W \tag{15–1}$$

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where Q is the net heat *added* to the system and W is the net work done *by* the system. We must be careful and consistent in following the sign conventions for Q and W. Because W in Eq. 15–1 is the work done *by* the system, then if work is done *on* the system, W will be negative and U will increase. Similarly, Q is positive for heat added to the system, so if heat leaves the system, Q is negative.

Heat added is +
Heat lost is −
Work on system is −
Work by system is +

Equation 15–1 is known as the **first law of thermodynamics**. It is one of the great laws of physics, and its validity rests on experiments (such as Joule's) to which no exceptions have been seen. Since Q and W represent energy transferred into or out of the system, the internal energy changes accordingly. Thus, the first law of thermodynamics is a great and broad statement of the *law of conservation of energy*.

First law of thermodynamics is conservation of energy

It is worth noting that the conservation of energy law was not formulated until the nineteenth century, for it depended on the interpretation of heat as a transfer of energy.

A given system at any moment is in a particular state and can be said to have a certain amount of internal energy, U. But a system does not "have" a certain amount of heat or work. Rather, when work is done on a system (such as compressing a gas), or when heat is added or removed from a system, the state of the system *changes*. Thus, work and heat are involved in *thermodynamic processes* that can change the system from one state to another; they are not characteristic of the state itself. Quantities which describe the state of a system, such as internal energy U, pressure P, volume V, temperature T, and mass m or number of moles n, are called **state variables**. Q and W are *not* state variables.

Internal energy is a property of the system; work and heat are not

EXAMPLE 15–1 **Using the first law.** 2500 J of heat is added to a system, and 1800 J of work is done on the system. What is the change in internal energy of the system?

APPROACH We apply the first law of thermodynamics, Eq. 15–1, to our system.

SOLUTION The heat added to the system is $Q = 2500\,\mathrm{J}$. The work W done *by* the system is $-1800\,\mathrm{J}$. Why the minus sign? Because 1800 J done *on* the system (as given) equals $-1800\,\mathrm{J}$ done *by* the system, and it is the latter we need to put in Eq. 15–1 by the sign conventions given above. Hence

$$\Delta U = 2500\,\mathrm{J} - (-1800\,\mathrm{J}) = 2500\,\mathrm{J} + 1800\,\mathrm{J} = 4300\,\mathrm{J}.$$

You may have intuitively thought that the 2500 J and the 1800 J would need to be added together, since both refer to energy added to the system. You would have been right.

NOTE We did this calculation in detail to emphasize the importance of keeping careful track of signs.

EXERCISE A What would be the internal energy change in Example 15–1 if 2500 J of heat is added to the system and 1800 J of work is done *by* the system (i.e., as output)?