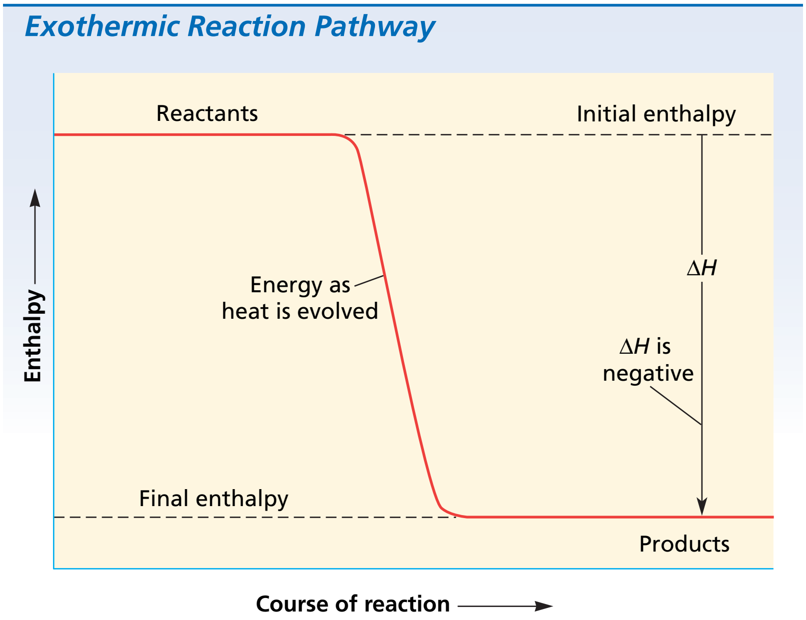
FIGURE 2 In an exothermic chemical reaction, the enthalpy change is negative, meaning energy is released from the system as heat.



$$2H_2O(g) + 483.6\ \text{kJ} \longrightarrow 2H_2(g) + O_2(g)$$

The physical states of reactants and products must always be included in thermochemical equations because they influence the overall amount of energy as heat gained or lost. For example, the energy needed for the decomposition of water would be greater than 483.6 kJ if we started with ice, because extra energy would be needed to melt the ice and to change the liquid into a vapor.

Thermochemical equations are usually written by designating the value of ΔH rather than writing the energy as a reactant or product. For an exothermic reaction, ΔH is always negative because the system loses energy. So, the thermochemical equation for the exothermic formation of 2 mol of gaseous water from its elements now has the following form.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g) \quad \Delta H = -483.6\ \text{kJ}$$

Figure 2 graphically shows the course of an exothermic reaction. The initial enthalpy of the reactants is greater than the final enthalpy of the products. This means energy as heat is evolved, or given off, during the reaction; this is described as a negative enthalpy change.

For an endothermic reaction, ΔH is always positive because the system gains energy. Thus, the endothermic decomposition of 2 mol of gaseous water has the following thermochemical equation.

$$2H_2O(g) \longrightarrow 2H_2(g) + O_2(g) \quad \Delta H = +483.6\ \text{kJ}$$

The course of an endothermic reaction is illustrated in **Figure 3.** Energy as heat is absorbed in this reaction, meaning that the initial enthalpy of