Meanwhile, metallic silver is oxidized at the anode according to the following half-reaction.

$$\overset{0}{Ag} \longrightarrow \overset{+1}{Ag^+} + e^-$$

In effect, silver is transferred from the anode to the cathode of the cell.

Rechargeable Cells

A rechargeable cell combines the oxidation-reduction chemistry of both voltaic cells and electrolytic cells. When a rechargeable cell converts chemical energy to electrical energy, it operates as a voltaic cell. But when the cell is recharged, it operates as an electrolytic cell, converting electrical energy to chemical energy.

The standard 12 V automobile battery, shown in **Figure 15,** is a set of six rechargeable cells. The anode in each cell is lead submerged in a solution of H_2SO_4. The anode half-reaction is described below.

$$Pb(s) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s) + 2e^-$$

At the cathode, PbO_2 is reduced according to the following equation.

$$PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \longrightarrow PbSO_4(s) + 2H_2O(l)$$

The net oxidation-reduction reaction for the discharge cycle is:

$$Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \longrightarrow 2PbSO_4(s) + 2H_2O(l)$$

A car's battery produces the electric energy needed to start its engine. Sulfuric acid, present as its ions, is consumed, and lead(II) sulfate accumulates as a white powder on the electrodes. Once the car is running, the half-reactions are reversed by a voltage produced by the alternator. The Pb, PbO_2, and H_2SO_4 are regenerated. A battery can be recharged as long as all reactants necessary for the electrolytic reaction are present, and all reactions are reversible.

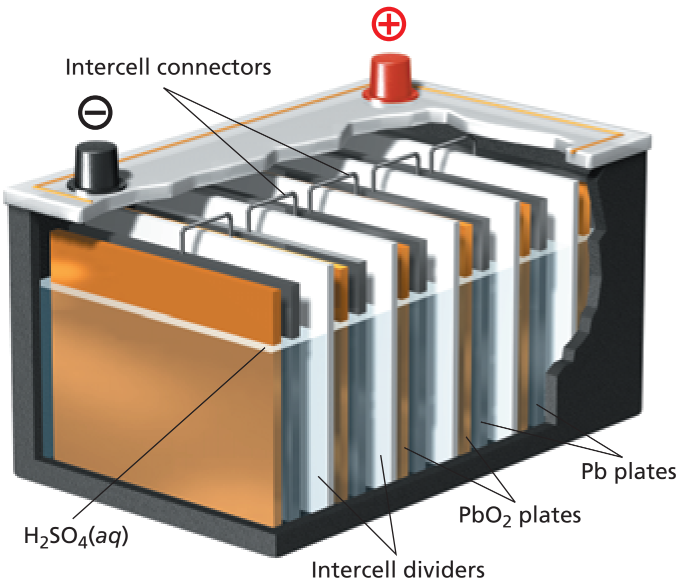

FIGURE 15 The rechargeable cells of a car battery produce electricity from reactions between lead(IV) oxide, lead, and sulfuric acid.