Reduction

Processes in which the oxidation state of an element decreases are **reduction** *processes.* Consider the behavior of chlorine in its reaction with sodium. Each chlorine atom accepts an electron and becomes a chloride ion. The oxidation state of chlorine decreases from 0 to –1 for the chloride ion (Rules 1 and 2, **Table 1**).

$$\overset{0}{Cl_2} + 2e^- \longrightarrow 2\overset{-1}{Cl^-}$$

A species that undergoes a decrease in oxidation state is **reduced.** The chlorine atom is reduced to the chloride ion.

FIGURE 3 Copper is oxidized and nitrogen dioxide is produced when this penny is placed in a concentrated nitric acid solution.

Oxidation and Reduction as a Process

Electrons are released in oxidation and acquired in reduction. Therefore, for oxidation to occur during a chemical reaction, reduction must also occur. Furthermore, the number of electrons produced in oxidation must equal the number of electrons acquired in reduction. This makes sense when you recall that electrons are negatively charged and that for charge to be conserved, the number of electrons lost must equal the number of electrons gained. You learned in Chapter 8 that mass is conserved in any chemical reaction. Therefore, the masses of the elements that undergo oxidation and reduction and the electrons that are exchanged are conserved.

A transfer of electrons causes changes in the oxidation states of one or more elements. *Any chemical process in which elements undergo changes in oxidation number is an* **oxidation-reduction reaction.** This name is often shortened to **redox reaction.** An example of a redox reaction can be seen in **Figure 3,** in which copper is being oxidized and NO_3^- from nitric acid is being reduced. *The part of the reaction involving oxidation or reduction alone can be written as a* **half-reaction.** The overall equation for a redox reaction is the sum of two half-reactions. Because the number of electrons involved is the same for oxidation and reduction, they cancel each other out and do not appear in the overall chemical equation. Equations for the reaction between nitric acid and copper illustrate the relationship between half-reactions and the overall redox reaction.

$$\overset{0}{Cu} \longrightarrow \overset{+2}{Cu^{2+}} + 2e^- \qquad \text{(oxidation half-reaction)}$$

$$2\overset{+5\,-2}{NO_3^-} + 2e^- + 4\overset{+1}{H^+} \longrightarrow 2\overset{+4\,-2}{NO_2} + 2\overset{+1\,-2}{H_2O} \qquad \text{(reduction half-reaction)}$$

$$\overset{0}{Cu} + 2\overset{+5}{NO_3^-} + 4H^+ \longrightarrow \overset{+2}{Cu^{2+}} + 2\overset{+4}{NO_2} + 2H_2O \qquad \text{(redox reaction)}$$

Notice that electrons lost in oxidation appear on the product side of the oxidation half-reaction. Electrons are gained in reduction and

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