hydrogen atom is oxidized. Neither atom has totally lost or totally gained any electrons. Hydrogen has donated a share of its bonding electron to the chlorine but has not completely transferred that electron. The assignment of oxidation numbers allows an approximation of the electron distribution of a molecule. An element can have different oxidation numbers in different compounds. This difference in oxidation numbers can reveal the difference in electron distribution of the compounds.

Reactants and products in redox reactions are not limited to monatomic ions and uncombined elements. Elements in molecular compounds or polyatomic ions can also be oxidized and reduced if they have more than one nonzero oxidation state. An example of this is provided in the reaction between the copper penny and nitric acid in which the nitrate ion, NO_3^-, is converted to nitrogen dioxide, NO_2. Nitrogen is reduced in this reaction. Usually, we refer to the oxidation or reduction of the entire molecule or ion. Instead of saying that the nitrogen atom is reduced, we say the nitrate ion is reduced to nitrogen dioxide.

$$\cdots + \overset{+5}{NO_3^-} \longrightarrow \overset{+4}{NO_2} + \cdots$$

SECTION REVIEW

1. How are oxidation numbers assigned?
2. Label each of the following half-reactions as either an oxidation or a reduction half-reaction:
   - **a.** $\overset{0}{Br_2} + 2e^- \longrightarrow 2\overset{-1}{Br^-}$
   - **b.** $\overset{0}{Na} \longrightarrow \overset{+1}{Na^+} + e^-$
   - **c.** $2\overset{-1}{Cl^-} \longrightarrow \overset{0}{Cl_2} + 2e^-$
   - **d.** $\overset{0}{Cl_2} + 2e^- \longrightarrow 2\overset{-1}{Cl^-}$
   - **e.** $\overset{+1}{Na^+} + e^- \longrightarrow \overset{0}{Na}$
   - **f.** $\overset{0}{Fe} \longrightarrow \overset{+2}{Fe^{2+}} + 2e^-$
   - **g.** $\overset{+2}{Cu^{2+}} + 2e^- \longrightarrow \overset{0}{Cu}$
   - **h.** $\overset{+3}{Fe^{3+}} + e^- \longrightarrow \overset{+2}{Fe^{2+}}$
3. Which of the following equations represent redox reactions?
   - **a.** $2KNO_3(s) \longrightarrow 2KNO_2(s) + O_2(g)$
   - **b.** $H_2(g) + CuO(s) \longrightarrow Cu(s) + H_2O(l)$
   - **c.** $NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$
   - **d.** $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$
   - **e.** $SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$
4. For each redox equation identified in the previous question, determine which element is oxidized and which is reduced.

Critical Thinking

5. **ANALYZING INFORMATION** Use the following equations for the redox reaction between Al^{3+} and Na to answer the questions below.

$$3\overset{0}{Na} \longrightarrow 3\overset{+1}{Na^+} + 3e^- \qquad \text{(oxidation)}$$

$$\overset{+3}{Al^{3+}} + 3e^- \longrightarrow \overset{0}{Al} \qquad \text{(reduction)}$$

$$3\overset{0}{Na} + \overset{+3}{Al^{3+}} \longrightarrow 3\overset{+1}{Na^+} + \overset{0}{Al} \qquad \text{(redox reaction)}$$

   - **a.** Explain how this reaction illustrates that charge is conserved in a redox reaction.
   - **b.** Explain how this reaction illustrates that mass is conserved in a redox reaction.
   - **c.** Explain why electrons do not appear as reactants or products in the combined equation.