This ratio is already in its lowest terms. If it were not, it would need to be reduced. Multiply the oxidation half-reaction by 1 (it remains unchanged) and the reduction half-reaction by 8. The number of electrons lost now equals the number of electrons gained.

$$1\left(\overset{-2}{H_2S} + 4H_2O \longrightarrow \overset{+6}{SO_4^{2-}} + 10H^+ + 8e^-\right)$$

$$8\left(\overset{+5}{NO_3^-} + 2H^+ + e^- \longrightarrow \overset{+4}{NO_2} + H_2O\right)$$

6. ***Combine the half-reactions, and cancel out anything common to both sides of the equation.***

$$\overset{-2}{H_2S} + 4H_2O \longrightarrow \overset{+6}{SO_4^{2-}} + 10H^+ + 8e^-$$

$$8\overset{+5}{NO_3^-} + 16H^+ + 8e^- \longrightarrow 8\overset{+4}{NO_2} + 8H_2O$$

$$8\overset{+5}{NO_3^-} + \overset{6}{\cancel{16}}H^+ + \cancel{8e^-} + \overset{-2}{H_2S} + \cancel{4H_2O} \longrightarrow$$

$$8\overset{+4}{NO_2} + \overset{4}{\cancel{8}}H_2O + \overset{+6}{SO_4^{2-}} + \cancel{10}H^+ + \cancel{8e^-}$$

Each side of the above equation has $10H^+$, $8e^-$, and $4H_2O$. These cancel each other out and do not appear in the balanced equation.

$$8\overset{+5}{NO_3^-} + \overset{-2}{H_2S} + 6H^+ \longrightarrow 8\overset{+4}{NO_2} + 4H_2O + \overset{+6}{SO_4^{2-}}$$

7. ***Combine ions to form the compounds shown in the original formula equation. Check to ensure that all other ions balance.*** The NO_3^- ion appeared as nitric acid in the original equation. There are only 6 hydrogen ions to pair with the 8 nitrate ions. Therefore, 2 hydrogen ions must be added to complete this formula. If 2 hydrogen ions are added to the left side of the equation, 2 hydrogen ions must also be added to the right side of the equation.

$$8HNO_3 + H_2S \longrightarrow 8NO_2 + 4H_2O + SO_4^{2-} + 2H^+$$

The sulfate ion appeared as sulfuric acid in the original equation. The hydrogen ions added to the right side are used to complete the formula for sulfuric acid.

$$8HNO_3 + H_2S \longrightarrow 8NO_2 + 4H_2O + H_2SO_4$$

A final check must be made to ensure that all elements are correctly balanced.

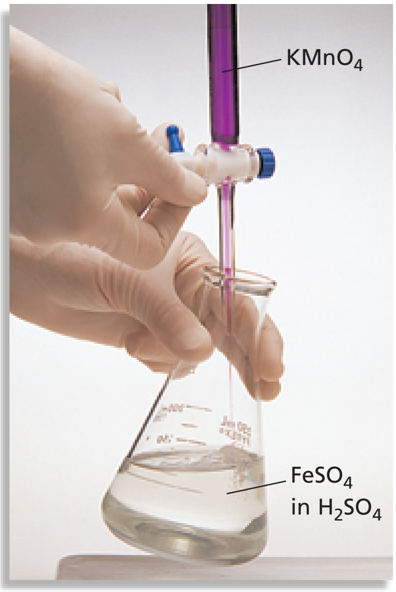

FIGURE 4 As a $KMnO_4$ solution is titrated into an acidic solution of $FeSO_4$, deep purple MnO_4^- ions are reduced to colorless Mn^{2+} ions. When all Fe^{2+} ions are oxidized, MnO_4^- ions are no longer reduced to colorless Mn^{2+} ions. Thus, the first faint appearance of the MnO_4^- color indicates the end point of the titration.

SAMPLE PROBLEM A

For more help, go to the Math Tutor at the end of this chapter.

Write a balanced equation for the reaction shown in Figure 4. A deep purple solution of potassium permanganate is titrated with a colorless solution of iron(II) sulfate and sulfuric acid. The products are iron(III) sulfate, manganese(II) sulfate, potassium sulfate, and water—all of which are colorless.