Name Key

$$\overset{0}{2Al_{(s)}} + \overset{+2}{3CuSO_{4(aq)}} \rightarrow \overset{+3}{Al_2(SO_4)_{3(aq)}} + \overset{0}{3Cu_{(s)}}$$

1. Write the oxidation half-reaction for the equation above.

   $\overset{0}{Al} \rightarrow \overset{+3}{Al} + 3e^-$ or $2Al \rightarrow 2Al^{+3} + 6e^-$

2. Write the reduction half-reaction for the equation above.

   $\overset{+2}{Cu} + 2e^- \rightarrow \overset{0}{Cu}$ or $3Cu^{+2} + 6e^- \rightarrow 3Cu$

$$\overset{+3\ -2}{Fe_2O_3} + \overset{0}{2Al} \rightarrow \overset{+3\ -2}{Al_2O_3} + \overset{0}{2Fe}$$

3. Write the oxidation half-reaction for the equation above.

   $\overset{0}{Al} \rightarrow \overset{+3}{Al} + 3e^-$ or $2Al \rightarrow 2Al^{+3} + 6e^-$

4. Write the reduction half-reaction for the equation above.

   $\overset{+3}{Fe} + 3e^- \rightarrow Fe$ or $2Fe^{+3} + 6e^- \rightarrow 2Fe$

$$\overset{0}{Zn_{(s)}} + \overset{+2\ -1}{SnCl_{2(aq)}} \rightarrow \overset{+2\ -1}{ZnCl_{2(aq)}} + \overset{0}{Sn_{(s)}}$$

5. Write the oxidation half-reaction for the equation above.

   $Zn \rightarrow Zn^{+2} + 2e^-$

6. Write the reduction half-reaction for the equation above.

   $Sn^{+2} + 2e^- \rightarrow Sn$

Use this balanced equation to answer the next question.

$$5H_2O_2 + 2KMnO_4 + 3H_2SO_4 \rightarrow 5O_2 + 2MnSO_4 + 8H_2O + K_2SO_4$$

7. If it takes 24.00 mL of 0.025 M $KMnO_4$ to oxidize 2.00 mL of H_2O_2 solution, what is the % of H_2O_2 in the solution? (Assume the H_2O_2 solution has a density of 1.00 g/mL)

$$\frac{.025 \text{ mol } KMnO_4}{1000 \text{ mL solution}} \times 24.00 \text{ mL solution} = \frac{.0006 \text{ mol } KMnO_4}{} \left| \frac{5 \text{ mol } H_2O_2}{2 \text{ mol } KMnO_4} \right| \frac{34.02 \text{ g } H_2O_2}{1 \text{ mol } H_2O_2}$$

$$= \frac{.05103 \text{ g } H_2O_2}{2.00 \text{ g } H_2O_2 \text{ solution}} \times 100 = 2.55\% \ H_2O_2$$

An iron II-containing ionic compound is dissolved in distilled water and sulfuric acid. The sample is titrated with 0.025 M $KMnO_4$ until a light pink color persists. The balanced equation and the data are given below.

$$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$

Mass of iron(II)-containing sample	0.45 grams
Molarity of $KMnO_4$ solution	0.025 M
Initial buret reading $KMnO_4$	10.50
Final buret reading $KMnO_4$	19.65

8. What is the % iron in the compound?

$$\frac{.025 \text{ mol } MnO_4^-}{1000 \text{ ml solution}} \times 9.15 \text{ ml solution} = .00022875 \text{ mol } MnO_4^- \left| \frac{5 \text{ mol } Fe^{2+}}{1 \text{ mol } MnO_4^-} \right| \frac{55.85 \text{ g } Fe^{2+}}{1 \text{ mol } Fe^{2+}} =$$

$$\frac{.06387... \text{ g } Fe^{2+}}{.45 \text{ g sample}} \times 100 = 14\% \text{ Fe}$$

The following balanced equation can be used for the next two questions.

$$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$

9. What volume (in Liters) of 1.00 M $KMnO_4$ will be required to oxidize all the iron II ions in a 40.00 gram sample of $FeCl_2 \cdot 4H_2O$? mL $126.75 + 72.08 = 198.8$ g

$$\frac{55.85 \text{ g Fe}}{198.8 \text{ g } FeCl_2 \cdot 4H_2O} = .281 \times 40.00 \text{ g} = 11.237... \text{ g } Fe^{2+} \left| \frac{1 \text{ mol } Fe^{2+}}{55.85 \text{ g } Fe^{2+}} \right| \frac{1 \text{ mol } MnO_4^-}{5 \text{ mol } Fe^{2+}} \left| \frac{1000 \text{ mL solution}}{1.00 \text{ mol } MnO_4^-} \right. =$$

40.2 mL $KMnO_4$ solution

10. If 20.00 mL of **0.10 M** $KMnO_4$ are mixed with 20.00 mL of **1.00 M** $FeSO_4$, what will be the color of the resulting solution? clear **Explain!**

$$\frac{.1 \text{ mol}}{1000 \text{ mL}} \times 20 \text{ mL} = .002 \text{ mol } KMnO_4$$

$$\frac{1 \text{ mol}}{1000 \text{ mL}} \times 20 \text{ mL} = .02 \text{ mol } Fe^{2+}$$

We have 10X more moles of Fe^{2+}. The reaction requires 5X more moles of Fe^{2+}, so we have more than enough Fe^{2+} to reduce all the MnO_4^- to the Mn^{2+} state.