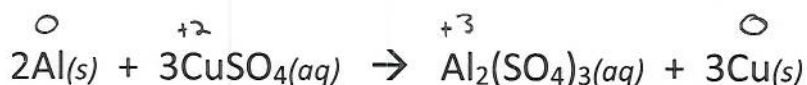
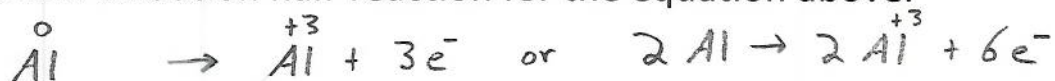


Name \_\_\_\_\_

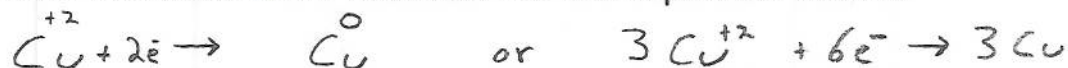
Key



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



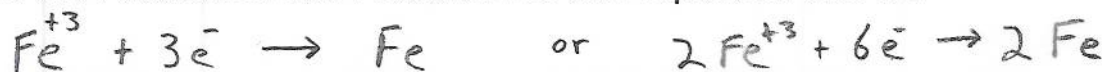
2. Write the reduction half-reaction for the equation above.



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



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



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



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



Use this balanced equation to answer the next question.



7. If it takes 24.00 mL of 0.025 M
- $\text{KMnO}_4$
- to oxidize 2.00 mL of
- $\text{H}_2\text{O}_2$
- solution, what is the % of
- $\text{H}_2\text{O}_2$
- in the solution? (Assume the
- $\text{H}_2\text{O}_2$
- solution has a density of 1.00 g/mL)

$$\frac{.025 \text{ mol KMnO}_4}{1000 \text{ mL solution}} \times 24.00 \text{ mL solution} = .0006 \text{ mol KMnO}_4$$

5 mol $\text{H}_2\text{O}_2$	34.02 g $\text{H}_2\text{O}_2$
2 mol $\text{KMnO}_4$	1 mol $\text{H}_2\text{O}_2$

$$= \frac{.05103 \text{ g H}_2\text{O}_2}{2.00 \text{ g H}_2\text{O}_2 \text{ solution}} \times 100 = 2.55\% \text{ H}_2\text{O}_2$$

An iron II-containing ionic compound is dissolved in distilled water and sulfuric acid. The sample is titrated with 0.025 M  $\text{KMnO}_4$  until a light pink color persists. The balanced equation and the data are given below.



Mass of iron(II)-containing sample	0.45 grams
Molarity of $\text{KMnO}_4$ solution	0.025 M
Initial buret reading $\text{KMnO}_4$	10.50
Final buret reading $\text{KMnO}_4$	19.65

8. What is the % iron in the compound?

$$\frac{0.025 \text{ mol MnO}_4^-}{1000 \text{ ml solution}} \times 9.15 \text{ ml solution} = 0.00022875 \text{ mol MnO}_4^-$$

$$\frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol MnO}_4^-} \times \frac{55.85 \text{ g Fe}^{2+}}{1 \text{ mol Fe}^{2+}} = 0.6387 \dots \text{ g Fe}^{2+}$$

$$\frac{0.6387 \dots \text{ g Fe}^{2+}}{0.45 \text{ g sample}} \times 100 = 14\% \text{ Fe}$$

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



9. What volume (in Liters) of 1.00 M  $\text{KMnO}_4$  will be required to oxidize all the iron II ions in a 40.00 gram sample of  $\text{FeCl}_2 \cdot 4\text{H}_2\text{O}$ ?  $126.75 + 72.08 = 198.8 \text{ g}$

$$\frac{55.85 \text{ g Fe}}{198.8 \text{ g FeCl}_2 \cdot 4\text{H}_2\text{O}} = 0.281 \times 40.00 \text{ g} = 11.237 \dots \text{ g Fe}^{2+}$$

$$\frac{1 \text{ mol Fe}^{2+}}{55.85 \text{ g Fe}^{2+}} \times \frac{1 \text{ mol MnO}_4^-}{5 \text{ mol Fe}^{2+}} \times \frac{1000 \text{ mL solution}}{1.00 \text{ mol MnO}_4^-} = 40.2 \text{ mL KMnO}_4 \text{ solution}$$

10. If 20.00 mL of 0.10 M  $\text{KMnO}_4$  are mixed with 20.00 mL of 1.00 M  $\text{FeSO}_4$ , what will be the color of the resulting solution? clear Explain!

$$\frac{0.1 \text{ mol}}{1000 \text{ mL}} \times 20 \text{ mL} = 0.002 \text{ mol KMnO}_4$$

$$\frac{1 \text{ mol}}{1000 \text{ mL}} \times 20 \text{ mL} = 0.02 \text{ mol Fe}^{2+}$$

We have 10X more moles of  $\text{Fe}^{2+}$ . The reaction requires 5X more moles of  $\text{Fe}^{2+}$ , so we have more than enough  $\text{Fe}^{2+}$  to reduce all the  $\text{MnO}_4^-$  to the  $\text{Mn}^{2+}$  state.