

Unit 4 – Chapter 4 Pre-Lab Questions:

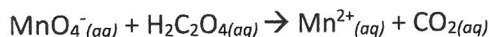
Name KEY

Permanganate Determination of the Iron Sample

Period _____

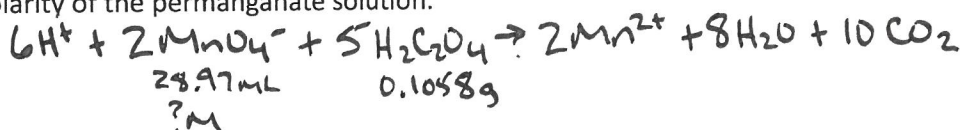
Redox prelab problems:

- 1) A solution of permanganate was standardized by titration with oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$). It required 28.97 mL of the permanganate solution to react completely with 0.1058 grams of oxalic acid. The unbalanced equation for the reaction is:



Calculate the molarity of the permanganate solution.

$$\begin{array}{l} 2 \times 1 = 2 \\ 2 \times 12 = 24 \\ 4 \times 16 = 64 \\ \hline 90 \text{ g H}_2\text{C}_2\text{O}_4 \end{array}$$

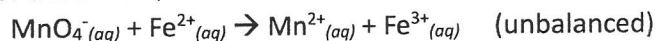


$$0.1058 \text{ g H}_2\text{C}_2\text{O}_4 \left| \frac{1 \text{ mol H}_2\text{C}_2\text{O}_4}{90 \text{ g H}_2\text{C}_2\text{O}_4} \right| \frac{2 \text{ mol MnO}_4^-}{5 \text{ mol H}_2\text{C}_2\text{O}_4} = 0.00047 \text{ mol MnO}_4^-$$

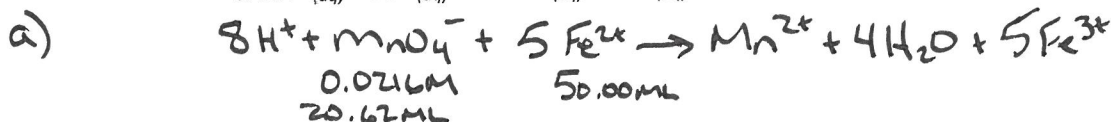
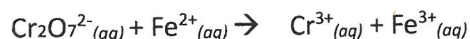
$$M = \frac{\text{mol Solute}}{\text{L Sol'n}}$$

$$M = \frac{0.00047 \text{ mol MnO}_4^-}{0.02897 \text{ L}} = \boxed{0.01622 \text{ M MnO}_4^-}$$

- 2) A 50.00 mL sample of solution containing Fe^{2+} ions is titrated with a 0.0216 M KMnO_4 solution. It required 20.62 mL of the KMnO_4 solution to oxidize all the Fe^{2+} ions to Fe^{3+} ions by the reaction:

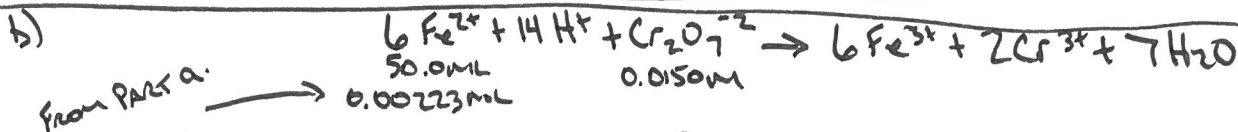


- a. What was the concentration of the Fe^{2+} ions in the sample solution?
b. What volume of 0.0150 M $\text{K}_2\text{Cr}_2\text{O}_7$ solution would it take to do the same titration? The reaction is:



$$0.0216 \text{ M} = \frac{X \text{ mol MnO}_4^-}{0.02062 \text{ L}} \rightarrow X = 0.000445 \text{ mol KMnO}_4 \left| \frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol KMnO}_4} \right| = 0.00223 \text{ mol Fe}^{2+}$$

$$\frac{0.00223 \text{ mol Fe}^{2+}}{0.0500 \text{ L}} = \boxed{0.0446 \text{ M Fe}^{2+}}$$

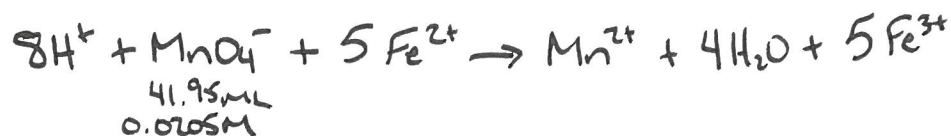


From Part a.

$$0.00223 \text{ mol Fe}^{2+} \left| \frac{1 \text{ mol Cr}_2\text{O}_7^{2-}}{6 \text{ mol Fe}^{2+}} \right| = 0.000372 \text{ mol Cr}_2\text{O}_7^{2-}$$

$$0.0150 \text{ M} = \frac{0.000372 \text{ mol Cr}_2\text{O}_7^{2-}}{X \text{ L Sol'n}} \quad \boxed{X = 24.8 \text{ mL}}$$

- 3) The iron content of iron ore can be determined by titration with standard KMnO_4 solution. The iron ore is dissolved in HCl and all the iron is reduced to Fe^{2+} ions. This solution is then titrated with KMnO_4 solution, producing Fe^{3+} and Mn^{2+} ions in acidic solution. If it required 41.95 mL of 0.0205 M KMnO_4 to titrate a solution made from 0.6128 grams of iron ore, what is the mass percent of iron in the iron ore?



$$\frac{0.0205 \text{ mol MnO}_4^-}{1 \text{ L SOLN}} \times \frac{0.04195 \text{ L}}{1} = 0.000860 \text{ mol KMnO}_4 \times \frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol KMnO}_4} = 0.00430 \text{ mol Fe}^{2+}$$

$$0.00430 \text{ mol Fe}^{2+} \times \frac{56 \text{ g Fe}^{2+}}{1 \text{ mol Fe}^{2+}} = 0.24 \text{ g Fe}^{2+}$$

$$\frac{0.24 \text{ g Fe}^{2+}}{0.6128 \text{ g ORE}} \times 100\% = \boxed{39.2\% \text{ Fe in ore}}$$