Name: _____ Date: _____

Redox Titration – Potassium Permanganate with an Iron(II) Salt

Objective:

To use a standardized potassium permanganate solution to determine the percentage of iron in an unknown sample.

Abstract:

Oxidation numbers describe the number of electrons an element will gain or lose during a reaction. Each atom in an equation can be assigned an oxidation number according to certain rules. If the oxidation number of an atom increases as you go from the left side to the right side if an equation, oxidation has occurred (electrons have been lost); if the oxidation number decreases, reduction has occurred (electrons have been gained).

$$MnO_4^- + 8\ H^+ + 5e^- \rightarrow Mn^{+2} + 4\ H_2O \qquad (1)$$

In the above reduction half-reaction, manganese has undergone a decrease in oxidation number from +7 to +2. In the following oxidation half-reaction, each iron atom has undergone an increase in oxidation number from +2 to +3.

$$Fe^{+2} \rightarrow Fe^{+3} + e^- \qquad (2)$$

Oxidation must occur along with reduction. The number of electrons lost and gained in the half-reactions must be equal. The overall redox reaction becomes:

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

In this experiment the density of a potassium permanganate solution will be used to stoichiometrically determine the mass of iron in a compound. From this, the percentage of iron can be determined.

Equipment:

250 mL Erlenmeyer flask

buret

ring stand

buret clamps

balance

100 mL graduated cylinder

Materials:

iron(II) compound

$KMnO_4$ solution

concentrated H_2SO_4

concentrated H_3PO_4

Procedure:

1. Rinse a buret thoroughly with distilled water. Rinse the buret carefully once with a 5 mL portion of $KMnO_4$ solution. After these rinses, fill the buret with the $KMnO_4$ solution. Be sure that the tip of the stopcock is filled.

2. Mass a 0.5 g sample of iron(II) unknown on a balance to the nearest 0.0001 g. Record the mass. Place the sample in a 250-mL Erlenmeyer flask. Dissolve the sample in 75 mL of distilled water. Add 5 mL of concentrated H_2SO_4, and 3 mL concentrated H_3PO_4 (85%) to the flask. Add a magnetic stir bar to the flask. The Fe^{+3} that will be produced during the titration forms a colorless complex with the PO_4^{-3} ions. This simplifies the detection of the endpoint.

3. Record the initial reading on the buret (read the bottom of the meniscus at eye-level) to the nearest 0.01 mL. Start to add the $KMnO_4$ solution. When the violet color of the MnO_4^- ion in the reaction does not disappear quickly, add the solution slowly. Towards the end of the titration, the solution should be added one drop at a time. When a faint pink color persists for 30 seconds, with constant swirling, the endpoint has been reached. A white piece of paper under the Erlenmeyer flask will aid in detecting color changes.

4. Record the level of $KMnO_4$ solution in the buret to 0.01 mL.

5. Repeat procedure for four total trials, and as desired, up to ten trials.

Data: $d_{solution} = 3.2251$ g/L

	Trial 1	Trial 2	Trial 3	Trial 4
Initial Vol $KMnO_4$ (mL)	2.45	15.67	30.45	3.54
Final Vol $KMnO_4$ (mL)	15.67	30.45	47.34	18.03
Mass Fe compound (g)	0.4987	0.5056	0.5134	0.5045

Calculations:

For each trial:

1. Find the volume of potassium permanganate used in each trial. Use this to calculate the mass of potassium permanganate used in each trial.

2. Stoichiometrically determine the mass of iron that reacted with the potassium permanganate for each trial. You are going to use the ratio of $MnO4^-$ to Fe^{+2} to determine this. Do not worry that KMnO4 is not given in the equation and only

MnO_4^- is shown. Each potassium permanganate particle yields one permanganate ion. The 1:1 ratio allows us to use thef MnO_4^- to Fe^{+2} ratio as shown.

3. Find the percentage of iron in the sample for each trial.

Once all of the percentages have been determined:

4. Report an average percentage after performing a Q-Test on the percentages.