

## Lab 4: How Can We Determine the Actual Percentage of H<sub>2</sub>O<sub>2</sub> in a Drugstore Bottle of Hydrogen Peroxide?

### Purpose:

To perform a titration to determine the concentration of a potassium permanganate solution. To use the standardized potassium permanganate solution in a titration of a H<sub>2</sub>O<sub>2</sub> solution. To calculate the percentage of H<sub>2</sub>O<sub>2</sub> in a bottle of drugstore hydrogen peroxide.

### Background:

Hydrogen peroxide, H<sub>2</sub>O<sub>2</sub>, is easily oxidized. It is used in commercial bleaching processes and in wastewater treatment plants as an environmentally friendly alternative to chlorine. Dilute solutions of H<sub>2</sub>O<sub>2</sub> are used to bleach hair and to clean wounds. Because it readily decomposes in the presence of light, heat, or metallic catalysts into water and oxygen, the quantity of a hydrogen peroxide solution must be checked regularly to maintain its effectiveness. The concentration of hydrogen peroxide can be analyzed by redox titration with potassium permanganate.

Titration is a method of volumetric analysis—the use of volume measurements to analyze the concentration of an unknown. The most common types of titrations are acid-base titrations, in which a solution of an acid, for example, is analyzed by measuring the amount of a standard base solution required to neutralize a known amount of the acid. A similar principle applies to redox titrations. If a solution contains a substance that can be oxidized, then the concentration of that substance can be analyzed by titrating it with a standard solution of a strong oxidizing agent.

Oxidation-reduction (redox) reactions involve a transfer of electrons between the species being oxidized and the species being reduced. The reactions are often balanced by separating the reaction components into half-reactions: *oxidation* (loss of electrons) and *reduction* (gain of electrons). In a redox reaction, the number of electrons lost by a species being oxidized is always equal to the number of electrons gained by the species being reduced. In the reaction being studied in this lab, solutions of hydrogen peroxide, H<sub>2</sub>O<sub>2</sub>, and potassium permanganate, KMnO<sub>4</sub>, will be combined in acidic solution. KMnO<sub>4</sub> will be used as the titrant to analyze the concentration of hydrogen peroxide in a drugstore bottle. The KMnO<sub>4</sub> solution is purple in color, as it is added to the hydrogen peroxide solution, it reacts with the H<sub>2</sub>O<sub>2</sub> to produce colorless Mn<sup>2+</sup>, thus the color will fade in the flask. When all the H<sub>2</sub>O<sub>2</sub> has been used up, the last drop of potassium permanganate that is added will keep its color. The endpoint of the titration is the point at which the last drop of KMnO<sub>4</sub> added to the solution causes it to turn pink.

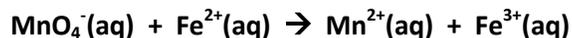
Solutions of KMnO<sub>4</sub> are not easily standardized solely by preparation, as solid KMnO<sub>4</sub> often contains impurities such as chlorides, sulfates, and nitrates. Before titrating with hydrogen peroxide, KMnO<sub>4</sub> will be standardized in acidic solution with a known concentration of iron (II) ions. In this oxidation-reduction reaction, manganese (II) ions and iron (III) ions are formed.

### Safety:

H<sub>2</sub>SO<sub>4</sub> is severely corrosive to eyes, skin, and other body tissues. Always add acid to water, never the reverse. KMnO<sub>4</sub> is a skin and eye irritant and a strong stain; it will stain skin and clothing. Goggles and aprons should be worn at all times. If solutions are spilled on skin, wash those areas immediately with copious amounts of water. Report all spills to the teacher. Wash hands thoroughly with soap and water after completion of the lab.

Pre-lab:

1. Balance the following redox reaction in acidic solution.



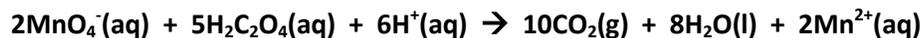
- What is the oxidation state of manganese before and after the reaction?
- What is the oxidation state of iron before and after the reaction?
- What species is being oxidized? What species is being reduced?
- How many electrons are transferred in this reaction?

2. Balance the following redox reaction in acidic solution.



- How many electrons are transferred in this reaction?
- What species is being oxidized? What species is being reduced?

3. A sample of oxalic acid,  $\text{H}_2\text{C}_2\text{O}_4$ , was analyzed using a standardized solution of  $\text{KMnO}_4$ . 25.0 mL of oxalic acid is titrated. 12.30 mL of a 0.0226 M  $\text{KMnO}_4$  was added to the sample when a faint pink color was observed, indicating the end point of the titration (meaning just enough  $\text{MnO}_4^-$  has been added to react with all 25.0 mL of oxalic acid). The balanced equation for this reaction is shown below:



- What is the ratio of  $\text{MnO}_4^-$  ions to  $\text{H}_2\text{C}_2\text{O}_4$  molecules in this reaction?
- How many moles of  $\text{MnO}_4^-$  ions reacted with the given amount of oxalic acid?
- How many moles of  $\text{H}_2\text{C}_2\text{O}_4$  were present?
- What was the molarity of the oxalic acid solution?
- If the density of the oxalic acid solution was 1.00 g/mL, what was the percentage by mass of oxalic acid in the solution?

4. Go to this website:

<http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/flashfiles/redoxNew/redox.html>

- Using the simulation, titrate  $\text{Fe}^{2+}$  with  $\text{KMnO}_4$ . Click 1. Select the Reaction. Choose  $\text{KMnO}_4$ .
- The buret will fill up with  $\text{KMnO}_4$  solution and the flask with  $\text{Fe}^{2+}$ .
- Use the slider to add large amounts of  $\text{KMnO}_4$ . Pay attention to the color of the solution in the flask. As you get closer to the end point, you will see pink in the flask.
- When you think you are getting close to the end point, add  $\text{KMnO}_4$  dropwise.
- When the solution in the flask stays pink, you have reached the end point. Be careful not to over titrate! If you over titrate, click "Repeat" to start over.
- Using the volume and molarity of  $\text{KMnO}_4$  necessary to reach the end point, calculate the molarity of  $\text{Fe}^{2+}$ . Enter the  $\text{Fe}^{2+}$  molarity into the box and click "OK" to check your answer. If incorrect, try again.
- If you want to practice again with another  $\text{KMnO}_4$  molarity, click "Reset".

Procedure:

**Part 1: Standardization of  $\text{KMnO}_4$  Solution**

1. Clean the buret as instructed.
2. Fill buret with  $\text{KMnO}_4$ . Be sure to get rid of any air bubbles by opening the stopcock. Discard into waste container. Volume of  $\text{KMnO}_4$  needs to be below zero mark on buret before starting titration.
3. Record the initial volume of  $\text{KMnO}_4$  to the correct number of significant figures.
4. Place 10.0 mL of 0.100 M  $\text{FeSO}_4$  into a 250 mL Erlenmeyer flask.
5. Place 5.0 mL of 6 M  $\text{H}_2\text{SO}_4$  into the same Erlenmeyer flask as the  $\text{FeSO}_4$ .
6. Swirl the contents of the flask.
7. Position the flask under the buret and place a piece of white paper under the flask to make it easier to detect the endpoint.
8. Open the buret stopcock and allow 5 mL of  $\text{KMnO}_4$  to flow into the flask. Swirl the flask and observe the color changes in the solution.
9. Continue adding  $\text{KMnO}_4$  a little at a time. The purple color should disappear when the flask is swirled. When it takes a few seconds to get rid of the purple color slow down and add  $\text{KMnO}_4$  drop by drop. When the reaction is complete, the solution turns pink. Record the final volume of  $\text{KMnO}_4$  when the endpoint is reached.
10. Dump the contents of the flask into the waste beaker.
11. Rinse the flask with tap water then distilled water and repeat the titration starting at Step 3. Record initial and final volumes for trial 2. Dump the contents of the flask into the waste beaker.
12. Clean buret as instructed. Wash all other used glassware.
13. Wipe lab bench with sponge.
14. Wash your hands with plenty of soap and water.

Data:

Table 1: Titration of 10.0 mL 0.100 M  $\text{FeSO}_4$  with  $\text{KMnO}_4$

Trial	Initial Volume of $\text{KMnO}_4$	Final Volume of $\text{KMnO}_4$	Volume of $\text{KMnO}_4$ used
1			
2			

Calculations:

1. Using the balanced equation from prelab, determine the moles of  $\text{KMnO}_4$  solution needed to react with all of the 0.100 M  $\text{FeSO}_4$  solution for each trial.
2. Calculate the molarity of the  $\text{KMnO}_4$  solution for each trial. Record each trial into the class data table.
3. Calculate the class average molarity of the  $\text{KMnO}_4$  solution.

## Part 2: Determination of the H<sub>2</sub>O<sub>2</sub> Concentration in a Bottle of Hydrogen Peroxide

1. Clean and prep a buret with KMnO<sub>4</sub> as done in Part 1.
2. Record the initial volume of KMnO<sub>4</sub> in the buret.
3. Place 1.00 mL of the hydrogen peroxide solution into a 250 mL Erlenmeyer flask.
4. Add about 25 mL of distilled water to the flask.
5. Place 5.0 mL of 6 M H<sub>2</sub>SO<sub>4</sub> into the flask.
6. Swirl the contents of the flask.
7. Titrate the hydrogen peroxide solution with the KMnO<sub>4</sub> solution. Record the final volume of KMnO<sub>4</sub>.
8. Dump the contents of the flask into the waste beaker and rinse the flask.
9. Repeat Steps 2-8 two more times. You should conduct three trials total.
10. Clean buret as instructed. Wash all other used glassware.
11. Wipe lab bench with sponge.
12. Wash your hands with plenty of soap and water.

### Data:

Table 2: Titration of 1.0 mL H<sub>2</sub>O<sub>2</sub> with KMnO<sub>4</sub>


### Calculations:

4. Using the molarity of KMnO<sub>4</sub> determined in Part 1, determine the moles of KMnO<sub>4</sub> for each trial.
5. Using the balanced equation from prelab, determine the moles of H<sub>2</sub>O<sub>2</sub> for each trial.
6. Calculate the grams of H<sub>2</sub>O<sub>2</sub> for each trial.
7. Assume the density of the H<sub>2</sub>O<sub>2</sub> solution to be 1.00 g/mL, calculate the percentage of H<sub>2</sub>O<sub>2</sub> in solution for each trial.
8. Calculate the average percentage of H<sub>2</sub>O<sub>2</sub> in solution.

### Post-lab:

1. What is a titrant? What is an analyte?
2. In this lab, which substance was used as both a titrant and analyte?
3. Compare your calculated average percentage to the value given on the bottle. Assume the bottle percentage to be the accepted value, calculate the percent error.

### Summary:

Include a summary of the results (average molarity of KMnO<sub>4</sub>, percentage of H<sub>2</sub>O<sub>2</sub>, and percent error) of the lab. Discuss any data values that are outliers. Discuss possible sources of error in determining the percentage of H<sub>2</sub>O<sub>2</sub>. Discuss why a titration was chosen to determine the percentage of H<sub>2</sub>O<sub>2</sub>.