Analysis of Hydrogen Peroxide

A Redox Titration

Introduction

Hydrogen peroxide is regarded as an "environmentally friendly" alternative to chlorine for water purification and wastewater treatment. Because hydrogen peroxide decomposes in the presence of heat, light, or other catalysts, the quality 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.

Concepts

- Redox reaction
- Oxidizing and reducing agents
- Titration
- Half-reactions

Background

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.

The equation for an oxidation-reduction reaction can be balanced by assuming that it occurs via two separate half-reactions. In this experiment, potassium permanganate will be used as the titrant to analyze the concentration of hydrogen peroxide in a commercial antiseptic solution. The permanganate ion acts as an oxidizing agent--it causes the oxidation of hydrogen peroxide. The oxidation half-reaction shows that two electrons are lost per molecule of hydrogen peroxide that is oxidized to oxygen gas (Equation 1). The permanganate ion, in turn, is reduced from the +7 oxidation state in MnO4- to the +2 oxidation state in Mn2+. The reduction half-reaction shows a gain of five electrons (Equation 2).

$$H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq) + 2e^- \qquad \textit{Equation 1}$$

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

Purple

Experiment Overview

The purpose of this experiment is to analyze the percent hydrogen peroxide in a common "drugstore" solution by titrating it with potassium permanganate. Standard potassium permanganate solution will be added via buret to the hydrogen peroxide solution. As the dark purple solution is added, it will react with the hydrogen peroxide and the color will fade. When all of the hydrogen peroxide 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 potassium permanganate added to the solution causes it to turn pink.

Materials
Distilled or deionized water, 100 mL
Hydrogen peroxide, H_2O_2, commercial antiseptic solution, 3 mL
Potassium permanganate solution, $KMnO_4$, 0.025 M, 75 mL
Sulfuric acid solution, HzSO4' 6 M, 15 mL
Beaker, 100- or 150-mL
Erlenmeyer flasks, 125-mL, 2
Labels and/or markers
Pipet bulb
Wash bottle
Buret, 50-mL, and buret clamp
Graduated cylinder, 10- or 25-mL
Pipet, volumetric or serological, I-mL
Ring stand
Waste disposal beaker, 250 mL

Safety Precautions
Sulfuric acid solution is severely corrosive to eyes, skin, and other body tissues. Always add acid to water, never the reverse. Notify your teacher and clean up all acid spills immediately. Potassium permanganate solution is a skin and eye irritant and a strong stain-it will stain skin and clothing. Avoid contact of all chemicals with eyes and skin. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the lab.

Procedure
1. Obtain about 75 mL of potassium permanganate standard solution in a small beaker. Record the precise molarity of the solution in the data table.

2. Rinse a clean 50-mL buret first with dH_2O with two 5-mL portions of potassium permanganate solution.

3. Clamp the buret to a ring stand using a buret clamp and place a waste beaker under the buret.

4. Fill the buret with potassium permanganate solution until the liquid level is just above the zero mark.

5. Open the stopcock on the buret to allow any air bubbles to escape from the tip. Close the stopcock when the liquid level in the buret is between the 0- and 5-mL mark.

6. Record the precise level of the solution in the buret. This is the *initial volume* of the potassium permanganate solution for Trial 1. *Note:* Volumes are read from the top down in a buret. Always read from the bottom of the meniscus and remember to include the appropriate number of significant figures.

7. Using a volumetric or serological pipet, transfer 1.00 mL of the commercial hydrogen peroxide solution into a 125-mL Erlenmeyer flask.

8. Add about 25 mL of distilled or deionized water to the flask.

9. Measure 5 mL of 6M sulfuric acid into a graduated cylinder and carefully add the acid to the solution in the Erlenmeyer flask. Gently swirl the flask to mix the solution.

10. Position the flask under the buret so that the tip of the buret is within the flask but at least 2 cm above the liquid surface. Place a piece of **<u>white paper under the flask</u>** to make it easier to detect the endpoint.

11. Open the buret stopcock and allow 5-8 mL of the potassium permanganate solution to flow into the flask. Swirl the flask and observe the color changes in the solution.

12. Continue to add the potassium permanganate solution slowly, drop-by-drop, while swirling the flask. Use a wash bottle to rinse the sides of the flask with distilled water during the titration to ensure that all of the reactants mix thoroughly.

13. When a light pink color persists in the titrated solution while swirling the flask, the endpoint has been reached. Close the stopcock and record the *final volume* of the permanganate solution in the data table (Trial 1).

14. Subtract the initial volume of the permanganate solution from the final volume to obtain the volume of $KMnO_4$ added. Enter the answer in the data table.

15. Pour the titrated solution into a waste disposal beaker and rinse the flask with distilled water.

16. Repeat the titration (steps 6-15) two more times (Trials 2 and 3). Record all data in the data table.

17. Dispose of the solution in the waste beaker as directed by your instructor.

Name(s) **Date:**

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A Redox Titration – Report Sheet

Pre-Lab Questions

1. Combine the oxidation and reduction half-reactions for hydrogen peroxide and permanganate ion, respectively, and write the balanced chemical equation for the overall reaction between H_2O_2 and MnO_4^- in acid solution. *Hint:* The number of electrons transferred must "cancel out."

2. What is the mole ratio of hydrogen peroxide to permanganate ion in the balanced chemical equation determined in Question #1? How many moles of hydrogen peroxide will be oxidized by 0.0045 moles of potassium permanganate in acidic solution?

3. Review the procedure. Is it necessary to know the exact volume of: (a) Hydrogen peroxide solution added to the flask in step 7? (b) Water added to the flask in step 8? Why or why not?

Data Table

	Trial 1	Trial 2	Trial 3
Molarity of $KMnO_4$			
Initial Volume $KMnO_4$ solution (mL)			
Final Volume $KMnO_4$ solution (mL)			
Volume of $KMnO_4$ solution used (mL)			

Calculations

1. Calculate the average volume of permangate ion used.

2. Calculate the moles of permangate ion used.

3. Calculate the number of moles hydrogen peroxide titrated.

4. Calculate the number of grams of hydrogen peroxide titrated.

5. Assuming the density of the hydrogen peroxide solution to be 1.00 g/mL, calculate the percent hydrogen perioxide by mass in the solution.

Analysis

1. Compare the value you calculated with value given on the side of the bottle.

2. If an insufficient amount of acid is added in step nine, some of the permanganate ions will be reduced to MnO_2 instead of Mn^{2+}.
   a. How would this change the mole ration for the calculations in the titration reaction?

   b. How would this affect the volume of the Potassium permanganate needed to reach the end point?

   c. If reduction to Mn were occurring but not reflected in the calculations, would the calculated percent hydrogen peroxide be too high of too low as a result of this error?