How Can We Determine the Actual Percentage of H_2O_2 in a Drugstore Bottle of Hydrogen Peroxide?

Hydrogen peroxide, H_2O_2, is easily oxidized. Dilute solutions are used as a disinfectant, in commercial bleaching processes and in wastewater treatment plants as an environmentally friendly alternative to chlorine. It readily decomposes in the presence of light, heat, or metallic catalysts into water and oxygen. It is important to know the actual concentration of a solution of H_2O_2 as its effectiveness can decrease with smaller concentrations.

Interestingly, hydrogen peroxide is also one of the two chief chemicals in the defense system of the bombardier beetle, reacting with hydroquinone to discourage predators. A study published in *Nature* found that hydrogen peroxide also plays a role in the immune system. Scientists found that hydrogen peroxide inside of cells increased after tissues are damaged in zebra fish, which is thought to act as a signal to white blood cells to converge on the site and initiate the healing process. When the genes required to produce hydrogen peroxide were disabled, white blood cells did not accumulate at the site of damage. The experiments were conducted on fish; however, because fish are genetically similar to humans, the same process is speculated to occur in humans. The study in *Nature* suggested asthma sufferers have higher levels of hydrogen peroxide in their lungs than healthy people, which could explain why asthma sufferers have inappropriate levels of white blood cells in their lungs.[1]

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 two half-reactions: *oxidation* (loss of electrons) and *reduction* (gain of electrons). In a redox reaction, the number of electrons lost by the 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_2O_2, and potassium permanganate, $KMnO_4$, will be combined in acidic solution.

Deep purple in solution, the manganese species in $KMnO_4$ undergoes reduction very easily. In acidic solution, permanganate ions (MnO_4^-) from $KMnO_4$ reduce to nearly colorless Mn^{2+} ions. In the presence of permanganate ions in acidic solution, an aqueous solution of H_2O_2 will undergo oxidation to make oxygen gas and hydrogen ions.

A titration, as you recall, is a convenient method of learning more about a solution by reacting it with a second solution of known molar concentration. There are a number of ways to measure the progress of a titration. The method used in this experiment is called a potentiometric titration, in which the electric potential of a reaction is monitored. All acid-base titrations that are measured by a pH probe are potentiometric; thus, this method is not as unusual as it may seem.

In this experiment you will use $KMnO_4$ in a titration to determine the concentration of the H_2O_2 solution. Your sample of H_2O_2 will come from a bottle of ordinary, over-the-counter hydrogen peroxide purchased at a grocery or drug store. Examine the label carefully and record the brand name and the percentage.

OBJECTIVES

In this experiment, you will

- Conduct the potentiometric titration of the reaction between commercially available hydrogen peroxide and potassium permanganate.
- Measure the potential change of the reaction with an oxidation-reduction probe (optional).
- Determine the concentration of the hydrogen peroxide solution.

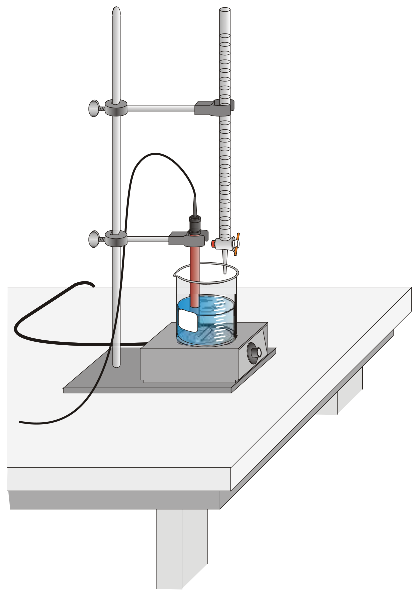
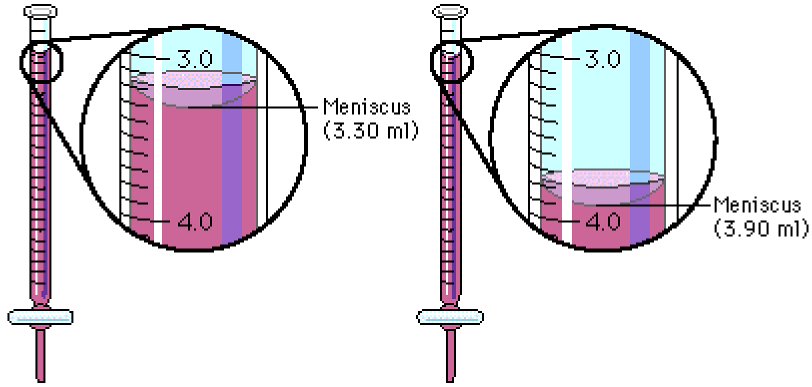


Figure 1

MATERIALS

Data Collection Device (optional)
Oxidation-Reduction Potential Sensor (optional)
ring stand
two 100 mL beakers
10 mL graduated cylinder
50 mL buret
magnetic stirrer and stirring bar *if available*
0.3% hydrogen peroxide, H_2O_2, solution
≈0.020 *M* potassium permanganate solution, $KMnO_4$
distilled water
utility clamp
buret clamp
stirring rod if no magnetic stirrer is available



How to properly read a buret!

PROCEDURE

1. Obtain and wear goggles.

2. a. Measure out precisely 10.00 mL of a ≈0.3% hydrogen peroxide solution from the dispensing buret your teacher has prepared into a 250 mL beaker. Should you miss the 10.00 mL mark, just record the volume of your sample *exactly* in the Data Table.

   b. Measure out approximately 10 mL of 4.5*M* sulfuric acid, H_2SO_4, solution from the dispensing buret your teacher has prepared into the beaker containing the ≈0.3% hydrogen peroxide solution.

   CAUTION: H_2SO_4 *is a **strong acid**, and should be handled with care.*

   c. Use a graduated cylinder to measure out approximately 25 mL of deionized water and add it to the beaker as well.

3. Place the beaker of H_2O_2 solution on a magnetic stirrer and add a stirring bar if available. If no magnetic stirrer is available, either stir the mixture with a stirring rod during the titration or swirl the mixture continuously.

4. OPTIONAL: Set up the data collection system. Connect an interface to the computer or handheld with the proper interface cable if necessary. If you are not using technology, then titrate to a pale pink endpoint.
   a. Connect an ORP Sensor the interface.
   b. Start the data collection program.
   c. Set up data collection for Events with Entry.

5. Set up a ring stand, a buret clamp, and a buret to conduct the titration (see Figure 1). Rinse and fill the 50 mL buret with the standardized $KMnO_4^-$ solution. Record its exact concentration in Data Table 2.

6. If you have elected to use the ORP sensor, place a utility clamp on the ring stand to hold the ORP Sensor in place during the titration. Position the ORP Sensor so that its tip is immersed in the H_2O_2 solution but does not interfere with the movement of the magnetic stir bar. Gently stir the beaker of solution.

7. Use your first trial to determine the volume of $KMnO_4$ needed to reach the equivalence point which is the largest increase in potential upon the addition of a very small amount of MnO_4^- solution or the volume at which a pale pink endpoint occurs. (Do not record this trial!) When you have completed this titration, dispose of the reaction mixture as directed. Rinse the ORP Sensor with distilled water in preparation for the first "real" trial.

8. Now that you have some idea of the volume of $KMnO_4$ needed to reach the equivalence point, repeat the titration carefully. When you have completed this titration, dispose of the reaction mixture as directed. Rinse the ORP Sensor with distilled water in preparation for the second "real" trial.

9. Record this value in your data table for Trial 1 and determine the *equivalence point*. If you used an ORP sensor, a good method of determining the precise equivalence point of the titration is to take the second derivative of the potential-volume data. To do this, plot Δ^2potential/Δvol^2 and examine the point where the curve crosses the zero line. The *x*-value at that point is the volume of MnO_4^- at the equivalence point. Record this value as the volume of MnO_4^- necessary to reach the equivalence point of the titration.

10. Repeat the necessary steps to conduct a second trial with a new sample of 0.3% H_2O_2 and acid solution.

11. When you conduct the second titration, carefully add the MnO_4^- solution drop by drop in the region near the equivalence point, so that you can precisely identify the equivalence point of the reaction.

12. Analyze the titration curve of your second trial to precisely determine the equivalence point of the reaction.

13. At the direction of your instructor, conduct a third trial. Print a copy of *each* titration curve.

DATA TABLE

	Trial 1	Trial 2	Trial 3
Volume of H_2O_2 solution (mL)			
Volume of MnO_4^- solution used (mL)			
Calculated number of moles of H_2O_2			

PRE-LAB QUESTIONS

1. What measuring devices are used to obtain precise measurements of volumes?

2. Write a balanced half-reaction for the reduction of permanganate ions in acidic solution. What are the oxidation states of manganese in this reaction?

3. Write a balanced half-reaction for the oxidation of hydrogen peroxide. What are the oxidation states of oxygen in this reaction?

4. How many moles of electrons are transferred in the balanced redox reaction? Justify your answer.

5. Write a balanced half-reaction for the oxidation of iron(II) ions.

6. Write a balanced half reaction for the reduction of permanganate ions by iron(II) ions. How many moles of electrons are transferred in this reaction?

7. Besides iron(II) ions and hydrogen peroxide, identify one or two other species that could be used to reduce permanganate ions.

8. Calculate the number of moles of MnO_4^- in 20.1 mL of 0.020 *M* potassium permanganate solution.

9. Calculate the number of moles of hydrogen peroxide needed to react with 20.5 mL of 0.020 *M* potassium permanganate solution.

10. The introduction states that the hydrogen peroxide solution sold in grocery stores is 3% (by mass). Assuming a density of 1.0 g/mL, calculate the molarity and mole fraction of 3.0 % H_2O_2.

POST-LAB QUESTIONS AND DATA ANALYSIS

1. Calculate the moles of MnO_4^- used to reach the equivalence point of the reaction for *each* trial.

2. Use the number of MnO_4^- moles to calculate the moles of H_2O_2 in the sample of solution for *each* trial.

3. Calculate the molar concentration of the H_2O_2 solution for *each* trial.

4. Calculate the Percent Error between your experimentally determined molarities for *each* trial and the molarity value for 3% hydrogen peroxide that you calculated as your answer to Pre-Lab Question 6. (Remember that you titrated 0.3% H_2O_2 solution samples and NOT 3% H_2O_2 solution samples.)

5. Why is hydrogen peroxide stored in a brown bottle?

6. A student fails to dry their 250 mL beaker between trials 2 and 3. What effect does this have on the calculated molarity of hydrogen peroxide for trial 3?

7. A student titrates three samples of 0.3% H_2O_2 solution made from a bottle of 3% H_2O_2 solution dated 10/31/2010. Would you expect the calculated molarities of the samples to be higher or lower than the value you calculated for your experimental samples? Justify your answer.