Experiment 17

Finding the Percent of H_2O_2 In Commercial Hydrogen Peroxide

Purpose

- Use the ideal gas law to determine the percentage of hydrogen peroxide in a commercially available hydrogen peroxide solution
- Observe how a catalyst affects a reaction

Background

Have you ever opened a container of milk after the expiration date and found that it had gone bad?

It is very easy to tell if milk has gone bad: it looks, smells, and tastes awful. The putrid odor lets you know a gas is being formed as it decomposes. Many items you purchase are not nearly as easy to tell if they have degraded. Some form a gas that is impossible to detect just by smelling it.

Hydrogen peroxide, a common household item that is used to clean minor cuts, is like this. The bottle you buy in the store says it contains 3% hydrogen peroxide, but it will very slowly decompose over time to form water and oxygen gas. We can't easily detect this, but we can use the ideal gas law and yeast to find this out how much has decomposed.

Since hydrogen peroxide forms oxygen gas when it decomposes, we can use the ideal gas law to check the percent hydrogen peroxide in a bottle of it purchased at the store. To find this out we need to take a small sample out of the bottle and accelerate its decomposition through using a catalyst

$$2\ H_2O_2 \xrightarrow{\text{catalyst}} 2\ H_2O + O_2$$

The ideal gas law is very valuable when dealing with gases since it establishes a relationship between temperature, pressure, volume, and amount of a gas
PV = nRT

In this equation:

- **P** is the gas pressure in atmospheres
- **V** is the volume of the gas in liters
- **n** is the number of moles of the gas
- **R** is the constant value of 0.08206 L·atm/mol·K
- **T** is for the temperature of the gas in Kelvin.

In this experiment, we will use yeast to accelerate the decomposition of the hydrogen peroxide into water and O_2 gas. Yeast contains the enzyme catalase, which is a catalyst for this reaction. You will add yeast to a known amount of hydrogen peroxide and quickly seal off the system so that the O_2 gas formed is collected in a graduated cylinder. After measuring the total volume of gas produced, its temperature, and the atmospheric pressure, the ideal gas law can then be used to calculate how many moles of O_2 gas is formed. We can do this by solving the ideal gas law equation for **n**.

$$n = \frac{PV}{RT}$$

Once the number of moles of O_2 gas is calculated, the percent of H_2O_2 present in the solution can be determined. To do this, you first need to calculate the theoretical number of moles of O_2 there would be if the solution was 100% hydrogen peroxide. This can be found by using the following equation

$$\textit{Theoretical moles } O_2 = H_2O_2 \textit{ used} \times H_2O_2 \textit{ density} \times \frac{1 \text{ mol } H_2O_2}{34.0 \text{ g } H_2O_2} \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O_2}$$

For this experiment:

- **mL H_2O_2** used is the volume of H_2O_2 you actually use (approximately 2-4 mL)
- **H_2O_2 density** is 1.02 g/mL
- **34.0 g H_2O_2 = 1 mol H_2O_2** (molar mass of H_2O_2)
- The molar ratio of O_2 : H_2O_2 from the reaction which is **1 mol O_2 / 2 mol H_2O_2**

The units in the entire equation cancel to give moles O_2

The percent hydrogen peroxide can be found by dividing the actual number of moles of oxygen you calculated by the theoretical number of moles of oxygen there would be if hydrogen peroxide were 100%. This number is then multiplied by 100%.

$$\% \, H_2O_2 = \frac{\text{Actual moles } O_2 \text{ (n)}}{\text{Theoretical moles } O_2} \times 100$$

This value can now be compared to the 3% hydrogen peroxide shown on the label to see if any decomposition has occurred.

EXAMPLE.

Consider the following examples involving another common chemical reaction which also produces oxygen gas.

The deposition reaction of potassium chlorate ($KClO_3$) produces potassium chloride (KCl) and oxygen gas (O_2)

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

If 32.5 mL of $KClO_3$ are used in the above reaction, how many moles of O_2 will be produced? The density of KClO3 is 2.34 g/mL

Solution

Convert the milliliters of $KClO_3$ to grams using its density and then to moles using its molar mass (122.55 grams/mole). Finally use the coefficients of the reaction to find moles of O_2 produced.

$$32.5 \, mL \;\; KClO3 \; x \frac{2.34 \; g \; KClO3}{1 \; mL \; KClO3} \frac{1 \; mole \; KClO3}{122.55 \; g \; KClO3} \; x \; \frac{3 \; mole \; O2}{2 \; mole \; KClO3} = \;\; 0.931 \; mole \; O_2$$

Equipment and Materials:

 15 mL dropper	 Plastic tubing	10 mL graduated cylinder 	250 mL breaker 
 Yeast	 Hydrogen peroxide	Weighting boat 	50 or 100 mL graduated cylinder 

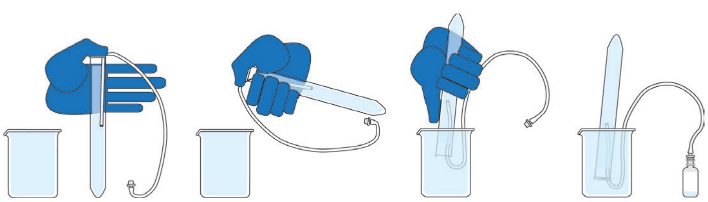
Procedure

1. Determine the current air temperature using the thermometer and record it in the Data Table.
2. Fill the 50-mL(or 100 mL) graduated cylinder completely with water.
3. Fill the 250-mL beaker with about 200 mL of water.
4. Attach the dropper top to one end of the clear tubing.
5. Insert the opposite end of the tubing into the 50-mL graduated cylinder.

Note: The next step will cause the water in the beaker to overflow; have paper towels ready to collect any spills.

6. Place thumb over opening of the 50-mL graduated cylinder and invert into the beaker. See Figure 2 below for an illustration.

Figure 2.



7. Record the initial volume of gas in the graduated cylinder in Data Table 1. If no gas is present, the volume will be 0. **If you have more than 15 mL of air in the graduated cylinder, you will need to repeat this step.**
8. Turn on the balance.
9. Place a weigh boat on the balance and tare it.
10. Weigh 0.1 gram of yeast into the weigh boat.
11. Pour the yeast into the empty 15-mL dropper bottle.
12. Measure 3 mL of hydrogen peroxide in the 10-mL graduated cylinder. Record the exact volume in the Data Table.
13. Prepare to quickly place the tubing with the dropper top on top of the dropper bottle.

Note: The top will need to be snapped in place immediately after the hydrogen peroxide is added.

14. Pour the hydrogen peroxide into the 15-mL bottle.
15. **Immediatel**y snap the dropper top with tubing onto the bottle.
16. Allow the reaction to proceed until no additional gas has been generated for at least one minute. Swirl the bottle to make sure all of the hydrogen peroxide has reacted. At this point, the reaction can be considered complete.
17. Record the final volume of gas in the Data Table.
18. Calculate the change in volume in mL and record in the Data Table.
19. Remove the top from the bottle, pour the contents down the sink, and rinse the bottle three times with water.
20. Repeat the activity varying the volume of hydrogen peroxide in **step 12** as listed in the table below.

Trial 2	Trial 3
2 mL H_2O_2	3.5 mL H_2O_2

Data Table:

	TRIAL 1	TRIAL 2	TRIAL 3
Temperature in Celsius			
Volume of H_2O_2 used (mL)			
Initial volume of air in graduated cylinder (mL)			
Final volume of air and O_2 in graduated cylinder after reaction (mL)			
Volume of O_2 collected (mL)			

Calculations

	TRIAL 1	TRIAL 2	TRIAL 3
Pressure *(given by instructor)*			
Pressure in atm			
Temperature in Kelvin			
Volume of O_2 in Liters			
R	0.08206 $\frac{\text{L atm}}{\text{mol K}}$	0.08206 $\frac{\text{L atm}}{\text{mol K}}$	0.08206 $\frac{\text{L atm}}{\text{mol K}}$
Actual moles of O_2 using PV = nRT			
Theorical moles of O_2			
percent hydrogen peroxide			

Sample Calculations For Trial 1:

Show Calculations for trial 1 below

The goal is to find the percentage of hydrogen peroxide in the solution!

This can be found by working through the following steps:

1. Convert the pressure given by your instructor into atmospheres

2. Convert the temperature from ^{o}C to Kelvin (K). Use the equation **$K = {}^{o}C + 273$**.

3. Convert the volume of oxygen from mL to liters (L).

4. You are now ready to solve for the number of moles of O_2. Be sure the units cancel so that you end up with only the moles of O_2 left.

Actual number of moles of O_2 (n) = ___________________ moles using PV = nRT

5. Calculate the **theoretical number of moles of O_2** there would be if the hydrogen peroxide were 100% and not an aqueous solution.

$$\textbf{Theoretical moles of } O_2 = H_2O_2 \text{ volume} * H_2O_2 \text{ density} * \frac{\text{mol } H_2O_2}{\text{g } H_2O_2} * \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O_2}$$

6. Find the percent hydrogen peroxide.

$$\% \ H_2O_2 = \frac{\text{Actual moles } O_2}{\text{Theoretical moles } O_2} * 100\% = \_\_\_\_\_\_\_\_\_\_ \ \%$$

Complete the calculations for the other trials. It is not necessary to show your work for the other trials however.

Name ________________________

Post-Lab Questions

1. Was the calculated percentage of hydrogen peroxide close to the same as the percentage on the label?

2. How many moles of nitrogen gas (N_2) can be produced by the following reaction if 41 grams of Mg_3N_2 are used?

$$Mg_3N_2 \rightarrow 3\ Mg + N_2$$

Name ______________________

Pre-Laboratory Assignment for

Finding the Percent of H_2O_2 in Commercial Hydrogen Peroxide

1. In the experiment, what gas will be produced by the decomposition of hydrogen peroxide?

2. What is the purpose of using yeast in this experiment's reaction?

3.Identify what each variable stands for in the ideal gas law ($PV = nRT$)

P is the

V is the

n is the

R is the

T is the

4.What units do the variables in $PV = nRT$ need to be in?

P needs to be in the unit

V needs to be in the unit

n needs to be in the unit

R needs to be in the units

T needs to be in the unit

5.What is the molar mass of hydrogen peroxide (H_2O_2)?

6. How many moles of O_2 were produced in a decomposition reaction of H_2O_2 if the pressure was 0.980 atm, the temperature was 298K and the volume of O_2 gas collected was 0.0520L?

7. If you decomposed 10.2 mL of 100% H_2O_2, how many moles of O_2 could you theoretically obtain?