Lab Manual
Introductory Chemistry: *A Green Approach*
Version 4.1

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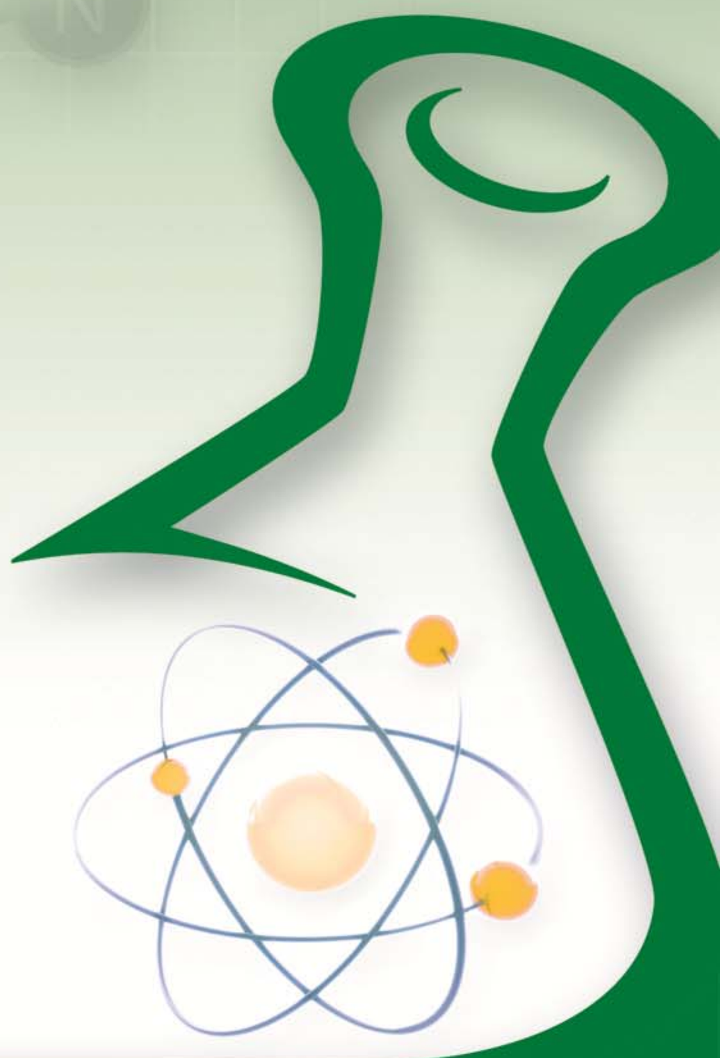
Introductory Chemistry



Lab 8:
Ideal Gas Law

Lab 8: Ideal Gas Law

Concepts to explore:

- 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
- Determine the decomposition rate of the hydrogen peroxide solution

Introduction

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. Since we breathe oxygen every second, we can't easily detect this. But we can use the ideal gas law and yeast to find this out!

Figure 1: Pressure gauges are found wherever monitoring the pressure of a gas is important—such as this fire extinguisher above. Simple gauges are commonly used to measure the pressure of air in automobile and bicycle tires.

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 0.0821 L·atm/mol·K
- **T** is for the temperature of the gas in Kelvin.

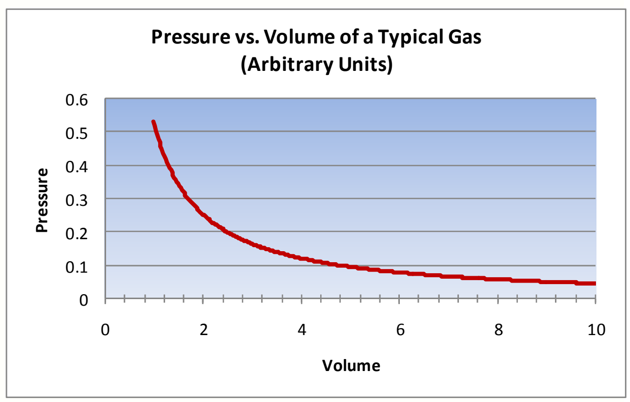


Figure 2: The relationship between pressure and volume of an ideal gas. With constant T and n, pressure decreases as volume increases. Can you verify this using the equation to the left?

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.

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

Lab 8: Ideal Gas Law

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 activated in warm water 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:

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

For this experiment:

- **mL H_2O_2 used** is the volume of H_2O_2 you actually use (approximately 5 mL).
- **H_2O_2 density** is 1.02 g/mL
- **1 mol H_2O_2 / 34.0 g H_2O_2** is the reciprocal (inverted fraction). of the molar mass of H_2O_2 . The molar mass of H_2O_2 is 34.0 g /mol, so this is equal to 1 mol H_2O_2 / 34.0 g H_2O_2.
- **1 mol O_2 / 2 mol H_2O_2** is used since the decomposition produces 1 mole O_2 from 2 moles of H_2O_2 .
- The units in the entire equation cancel to give moles of O_2.

Figure 3: **Carbonated beverages contain dissolved CO_2 at high pressure. When the container is opened, this pressure can create a powerful burst, such as with this sparkling wine bottle, or when your soda "explodes."**

The percent hydrogen peroxide can now be found. To do this, divide (**n**), the actual number of moles you calculated, by the **theoretical moles of O_2** there would be if the hydrogen peroxide were 100%. This number is then multiplied by 100%.

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

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

Pre-Lab Questions

1. What is it in yeast that aids in the decomposition of hydrogen peroxide?

2. List the ideal gas law and define each term with units.

3. How many moles of O_2 were produced in a decomposition reaction of H_2O_2 if the barometric pressure was 0.980 atm, the temperature was 298 K and the volume of O_2 gas collected was 0.0500 L?

4. If you decomposed 10.00 mL of 100% H_2O_2, how many moles of O_2 could you theoretically obtain?

Lab 8: Ideal Gas Law

Experiment: Finding Percent H_2O_2 with Yeast

Materials

Safety Equipment: Safety goggles, gloves

Yeast
10 mL Hydrogen peroxide
10 and 100 mL Graduated cylinders
Erlenmeyer flask
Stopper with hole
Rigid plastic tubing (3 in.)
Rubber band
2 Droppers (pipettes)
Stir rod
Thermometer
Warm water*
Large ring*
Flexible tubing (18 in.)
250 mL Beaker
600 mL Beaker
Stopwatch
Ring stand*
Distilled water*

*You must provide
*Optional Materials (not provided)

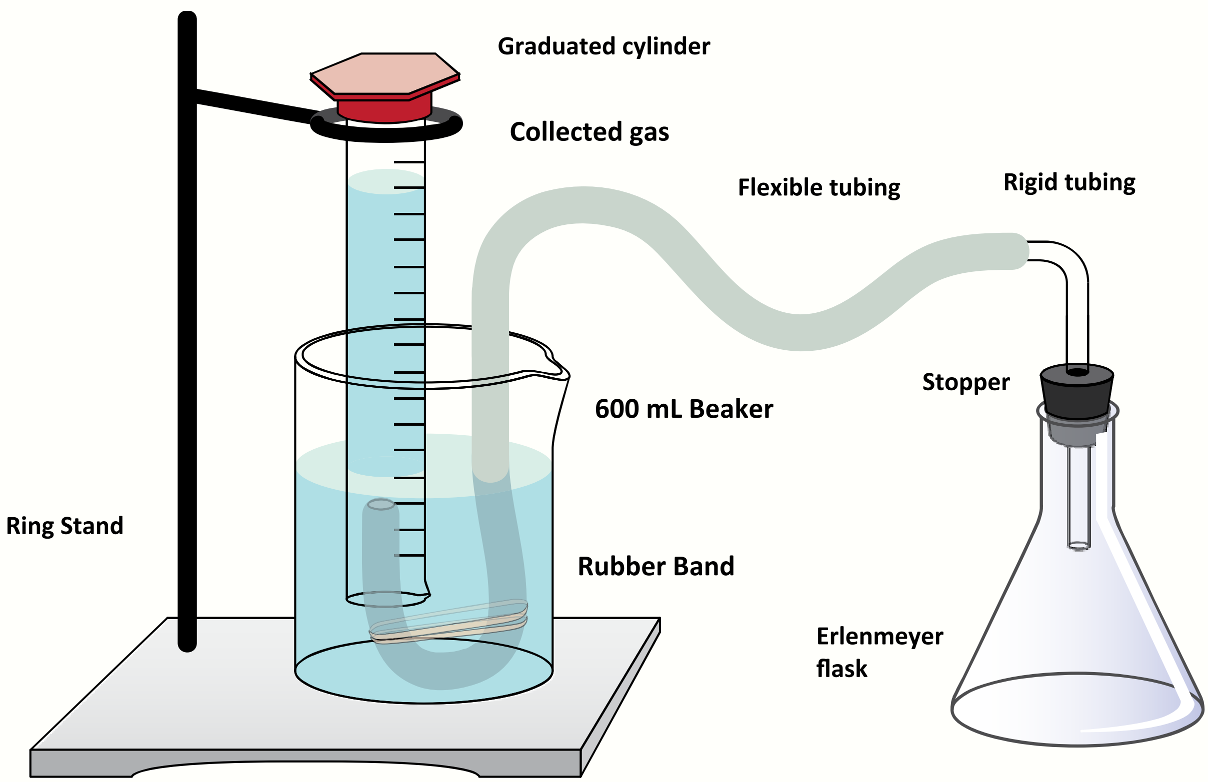


Figure 3: Gas Collection Apparatus (not to exact scale)

Procedure

1. Prepare the materials for the apparatus as shown in Figure 1. Insert the smaller rigid tubing into one end of the larger, flexible tubing. Insert the free end of the rigid tubing securely into the rubber stopper hole.

2. Bend the free end of the flexible tubing into a U shape, and use a rubber band to hold this shape in place. This will allow you to more easily insert this end of the flexible tubing into the inverted graduated cylinder. Make sure the tubing is not pinched and that gas can flow freely through it.

2. Fill the 600 mL beaker with 400 mL distilled water.

3. Fill the 100 mL graduated cylinder with distilled water slightly over the 100 mL mark.

4. Take the temperature of the water in the 600 mL beaker, and record it in the Data section. Also, determine the barometric pressure in the room, and record it in the Data section. **HINT:** The pressure in your region may be found online—if necessary, convert this value to mm Hg.

5. Mix 100 mL of warm water (45°C) and 1 packet of baker's yeast in a 250 mL beaker. This will activate the yeast from the dormant (dry) state. Be sure to mix well with a stir rod until the yeast is completely dissolved.

6. Use a 10 mL graduated cylinder and pipette to measure out 5.00 mL of hydrogen peroxide. Pour this hydrogen peroxide into the Erlenmeyer flask, and place the stopper with stopper tube over the top.

7. Clean the 10 mL graduated cylinder by rinsing it at least three times with distilled water. Dispose of the rinse down the drain.

8. Cover the opening of the graduated cylinder with two or three fingers and **quickly** turn it upside down into the 600 mL beaker already containing 400 mL of water. DO NOT remove your fingers from the opening until the graduated cylinder is fully submerged under the water. If the amount of trapped air exceeds 10 mL, refill the cylinder and try again.

9. Insert the U shaped side of the flexible tubing into the beaker, and carefully snake it into the submerged opening of the graduated cylinder. You want as little air as possible to be in the graduated cylinder.

10. Secure the graduated cylinder to the ring stand by sliding a ring under the submerged cylinder, then attaching the ring to the stand.

    OPTIONAL PROCEDURE: If your kit does not include a ring stand, you will hold the graduated cylinder in place while gas is collected. Make sure to keep the open end of the cylinder completely submerged to prevent additional gas from entering. Rest the graduated cylinder against the side of the beaker during experimental setup.

11. With the cylinder vertical, record the volume of air inside (the line at which the water reaches in the cylinder) in the Data section in Table 1.

12. Using the pipette, measure out 5.00 mL of yeast solution into the rinsed 10 mL graduated cylinder. **NOTE:** Do not immediately pour the yeast solution into the Erlenmeyer flask.

13. Prepare to place the stopper (still connected to the hose) on the Erlenmeyer flask. Reset the stopwatch.

14. Quickly pour the 5.0 mL of yeast solution into the Erlenmeyer flask. Immediately place the stopper securely in the opening of the Erlenmeyer flask by twisting it down into the flask gently.

15. Start timing the reaction with the stopwatch.

16. Swirl the Erlenmeyer flask to mix the two solutions together.

17. You will begin to see bubbles coming up into the 100 mL graduated cylinder. **HINT:** If gas bubbles are not immediately visible, make sure the stopper is on tight enough and the tubing is not leaking. You will need to start over after correcting any problems.

18. Continue to swirl the Erlenmeyer flask and let the reaction run until **no more** bubbles form to assure the reaction has gone to completion. This should take approximately 6-10 minutes. **HINT:** Catalase works best around the temperature of the human body. You can speed the reaction up by warming the Erlenmeyer flask with your hands.

19. Record the time when the reaction is finished in Table 2 of the Data section, along with the final volume of air in Table 1. Remember to read it at eye-level and measure from the bottom of the meniscus.

20. Pour all other liquids down the drain and clean the labware.

Lab 8: Ideal Gas Law

Data

Water temperature: ____________________ ^oC

Barometric Pressure: ____________________ mm Hg

Table 1: Volume data

Initial volume of air (mL)	Final volume of air after reaction (mL)	Volume of O_2 collected (Final volume - initial volume)

Table 2: Reaction time data

Time reaction started	Time reaction ended	Reaction time (s)

Calculations

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 temperature of the water from ^oC to Kelvin (K). Use the equation **K = ^{o}C + 273.** This will be your value for absolute **T** or the temperature in Kelvin.

$$\mathbf{T} = \underline{\hspace{3cm}}\ ^oC + 273 = \underline{\hspace{3cm}}\ K$$

2. If necessary, convert the barometric pressure in the room from mm Hg to atmospheres (atm).

   Divide the measured pressure from the Data section by 760 mm Hg. This will give you pressure (**P**) in atmospheres.

$$\mathbf{P} = \underline{\hspace{2.5cm}}\ \text{mm Hg} * \frac{1\ \text{atm}}{760\ \text{mm Hg}} = \underline{\hspace{2.5cm}}\ \text{atm}$$

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

$$\mathbf{V} = \underline{\hspace{4cm}} \text{ mL} * \frac{1\text{ L}}{1000\text{ mL}} = \underline{\hspace{4cm}} \text{ L}$$

4. Rearrange the ideal gas law to solve for **n.**

5. 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. Use the value for the constant R given:

R = 0.0821 L·atm/mol·K

***Actual* number of moles of O_2 (n) =** ______________________ moles

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

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

To use the above equation, calculate the following:

- **H_2O_2 volume** is the volume (mL) of hydrogen peroxide used: Volume = ____________ mL H_2O_2
- **H_2O_2 density** is known: Density = **1.02 g/mL**
- $\dfrac{\textbf{mol } \mathbf{H_2O_2}}{\textbf{g } \mathbf{H_2O_2}}$ is the reciprocal of the molar mass of H_2O_2. First write the molar mass of H_2O_2 then find the reciprocal.

Molar mass of H_2O_2 = ____________________ g H_2O_2/1 mol H_2O_2

Molar mass of H_2O_2 reciprocal = ____________________

Now you have all of the information needed to solve the equation for the theoretical moles of O_2. All you need to do is fill in the blanks and do the calculations.

Theoretical moles of O_2 = __________ * __________ * __________ * __________

Theoretical moles of O_2 = ____________________ mol

7. Find the percent hydrogen peroxide.

$$\% \ H_2O_2 = \frac{\text{Actual moles } O_2}{\text{Theoretical moles } O_2} * 100\% = \underline{\hspace{4cm}} \%$$

8. You can also easily determine the reaction rate. To do this, divide the total volume of oxygen collected by the total time of the reaction.

$$\textbf{Reaction rate} = \frac{\text{Volume } O_2 \text{ (mL)}}{\text{Reaction time (s)}} = \underline{\hspace{5cm}} \text{ mL/sec}$$

Post-Lab Questions

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

2. Considering that catalysts are not consumed in a reaction, how do you think increasing the amount of catalyst would affect the reaction rate for the decomposition of hydrogen peroxide?

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