**Abstract**

 The Decomposition of hydrogen peroxide can demonstrate Dalton’s law of partial pressure. Dalton’s Law calculates the total pressure in a given mixture. Hydrogen peroxide decomposes slowly; therefore, a catalyst was used to speed up the reaction. The average percent from 3 trials of hydrogen peroxide was found to be $2.538 \%$ and the percent of error was -15.4%.

**Introduction**

 Dalton’s Law of partial pressure can determine the pressure in a mixture. The Law states in order to find the total pressure of a mixture; each pressure must be determined independently. Using a catalyst such as potassium iodide can speed up the reaction to examine what happens. The procedure involves decomposing hydrogen peroxide producing oxygen and water.

$$2H\_{2}O\_{2}\left(aq\right) \vec{KI} 2H\_{2}O(l)+O\_{2}(g)$$

 The main purpose of this experiment is to find the percent of hydrogen peroxide in the reaction. The technique involves measuring approximately 5 grams of hydrogen peroxide solution into the Erlenmeyer flask. The 800 mL beaker was filled with approximately 800 mL of water. The graduated cylinder was filled with water and carefully placed over the beaker. Measuring the initial air inside the cylinder was important to find the final volume. A syringe containing nearly 3 mL of potassium iodide was connected to the flask to speed up the reaction. When the reaction is initiated the air produced from the reaction flows through a tube and displaces the water in the graduated cylinder. Dalton’s Law can calculate the pressure of oxygen since the total pressure was a combination of oxygen and water in the graduated cylinder.

$P\_{Total}=$ P$o\_{2}+P\_{H\_{2}o}$

 The mixture total pressure was calculated using the fact that it displaces 30.05 in Hg. By knowing the pressure of oxygen the ideal gas law can be used to find the number of moles of oxygen.

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

Knowing the number of moles of oxygen in the reaction and using the balanced equation above, the number of moles of hydrogen peroxide in the solution can be determined. Dividing by the average mass collected in the experiment and multiplying by the molar mass will yield the amount of hydrogen peroxide in grams. Multiplying by a hundred will give the percent of hydrogen peroxide in the solution. To determine the volume of oxygen gas  the equation blew was used.

$V\_{O2}=$ $V\_{final}-V\_{initial\_{ }}-3 mL$

Note that for first trial, instead of subtracting 3 mL, we used 2.5 mL instead. This was done due to the lost of the KI solution that was added.

**Data**

Table 1: Collected data

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
|  | Temperature C | Temp K | Initial Volume, mL | Final Volume, mL | Mass of H2O2 g | Vapor Pressure(torr) |
| Trial 1 | 24.1 | 297.25 | 9 | 59 | 5.17 | 22.5 |
| Trial 2 | 23.3 | 296.45 | 9 | 55 | 5.09 | 21.3 |
| Trial 3 | 23.1 | 296.25 | 5 | 60 | 5.05 | 21.2 |

Table 2 Data used to calculate n

|  |  |
| --- | --- |
| Pressure, torr | 763.3 |
| Molar mass of H2O2 | 34 |

 **Results**

Table 3:

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
|    | Volume of O2 mL | n for O2 moles of O2 | Moles of H2O2 Moles  | Initial H2O2 mass, g | Mass Percent, % | Percent Error% |
| Trial 1 | 0.04750 | 0.00190 | 0.00380 | 0.12908 | 0.02497 | -16.77710 |
| Trial 2 | 0.04300 | 0.00173 | 0.00345 | 0.11736 | 0.02306 | -23.14648 |
| Trial 2 | 0.05200 | 0.00209 | 0.00418 | 0.14203 | 0.02813 | -6.24883 |

$$Avarage ercent Error \left(\%\right) =-15.4\%$$

**Discussion**

 The experiment went smoothly, three trials were done for. This result may have been caused from air leak in the Erlenmeyer flask from the rubber stopper. Shown on table 1 are the collected data for the 3 trials. The total pressure and the pressure of the air were calculated from the given data. The total pressure given in mercury was converted to atmospheric pressure. Therefore, the pressure of oxygen was calculated using Dalton’s Law of partial pressure (Table 2). The number of moles of oxygen in the solution was calculated using the ideal gas law (Table 3). The number of moles of hydrogen peroxide was divided by the average mass used in the experiment. Multiplying by the molar mass and a hundred gave the percent of hydrogen peroxide. The final percentage of hydrogen peroxide was $2.538 \%$, which is kind of low due to the fact that the accepted value is 7 %. This yields $15.4\%$ error, which is acceptable, The experiment was done according to the manual, however, some sources of error may have occurred. The possible source of error is that the stockroom prepared the solution with a small initial mass of hydrogen peroxide. Thus, causing a small percent of decomposition of hydrogen peroxide. Also, error may have occurred caused by the measurements instruments, starting with making the reaction. Another source of error was when trying to mix the reaction we hit the temperature and change its position which may have affected us in our reading and uncertainty.

Appendix:

Calculation for all three trials.