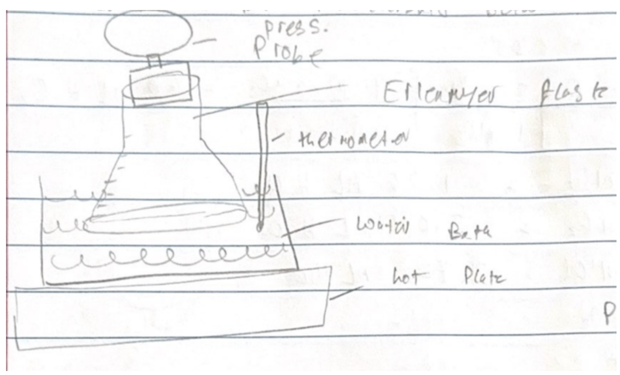
**Exploring Gas Laws**
10/9/2023
CHEM 1211K Laboratory

***Methods***

**Carrying out Part D: Relationship of Temperature with Pressure**

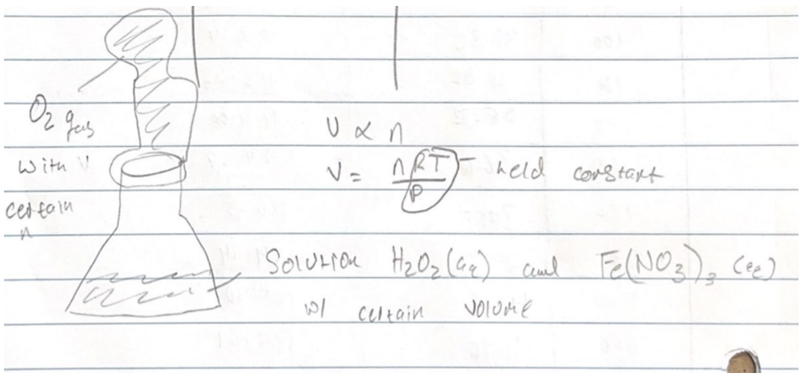
To obtain data to determine a relationship between temperature and pressure of a gas, we must hold the number of moles of such gas and the volume of the apparatus constant. In this experiment, we will be seeing how pressure is affected by temperature. Refer to the image below and the following steps for carrying out the experiment.




1. Obtain a 50 mL Erlenmeyer Flask, alcohol thermometer, 500 mL beaker filled halfway with tap water, hot plate, and a pressure probe.
2. Ensure the 50 mL Erlenmeyer flask is empty and close the top with a rubber stop with an inlet into the pressure probe.
3. Place the 50 mL Erlenmeyer Flask in the other beaker filled with water. Record the initial temperature and pressure of the water bath and Erlenmeyer flask, respectively.
4. Turn on the hot plate and record the pressure and temperature of the apparatus every minute until the overall temperature change is 10˚C in the water bath ($\Delta T = 10^\circ C$).

**Carrying out Part E: Relationship of Volume with Moles of Oxygen gas**

To carry out part E of this experiment, we need to determine a relationship between volume and number of moles of oxygen gas without varying the temperature of the gas nor the pressure. As such, this experiment will NOT keep the reaction insulated, and will let heat travel freely between the beaker and the surrounding environment. Additionally, the pressure will not be restricted; it will be able to keep constant pressure by varying volume. Refer to the image below while carrying out this experiment. We will be measuring, specifically, the volume of oxygen gas that comes out of the decomposition of hydrogen peroxide. We will be using excess Iron (III) Nitrate as a catalyst for this reaction.




1. First, determine a set amount of moles of oxygen gas you want. A range from .01 to .05 moles is acceptable. Use this theoretical amount to compute the volume of hydrogen peroxide needed for this experiment by referring to the chemical equation: $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$
2. Measure out this amount of Hydrogen Peroxide precisely and pour it into a 50 mL Erlenmeyer Flask.
3. Obtain a balloon and add 5 mL of Iron (III) Nitrate to it. The exact amount it not pertinent – just a rough amount.
4. Wrap the opening of the balloon around the opening of the Erlenmeyer flask making sure not to spill the Iron (III) Nitrate solution inside.
5. Once you have ensured the seal is tight, dump the Iron (III) Nitrate solution into the Hydrogen Peroxide at the bottom of the flask and let the reaction carry out. Oxygen gas should be produced as a result and should start expanding the balloon.
6. After you see that the reaction is no longer contributing oxygen gas to the balloon, carefully remove the balloon from the flask and tie a knot at the end.
7. Obtain a large jug with a known volume water initially inside. Submerge the balloon into the jug and measure the new volume the water rises to. By finding the difference between the initial and the final volumes, you determine the volume of oxygen gas produced by the reaction.

***Data and Results***

| Part D: Relationship of T vs P | | | | |
|---|---|---|---|---|
| Trial | Temperature (˚C) | Temperature (˚K) | Pressure (kPa) | Pressure (atm) |
| 1 | 26.57 | 299.74 | 101.06 | 0.9974 |
| 2 | 28.95 | 302.12 | 101.98 | 1.0065 |
| 3 | 29.7 | 302.87 | 102.29 | 1.0095 |
| 4 | 31.07 | 304.24 | 102.61 | 1.0127 |
| 5 | 32.32 | 305.49 | 103.14 | 1.0179 |
| 6 | 33.82 | 306.99 | 103.52 | 1.0217 |
| 7 | 36.63 | 309.8 | 104.37 | 1.0301 |
| 8 | 39.38 | 312.55 | 104.94 | 1.0357 |
| 9 | 40.82 | 313.99 | 105.1 | 1.0373 |
| 10 | 44.8 | 317.97 | 106.55 | 1.0516 |

Table 1: Data of Temperature with Pressure

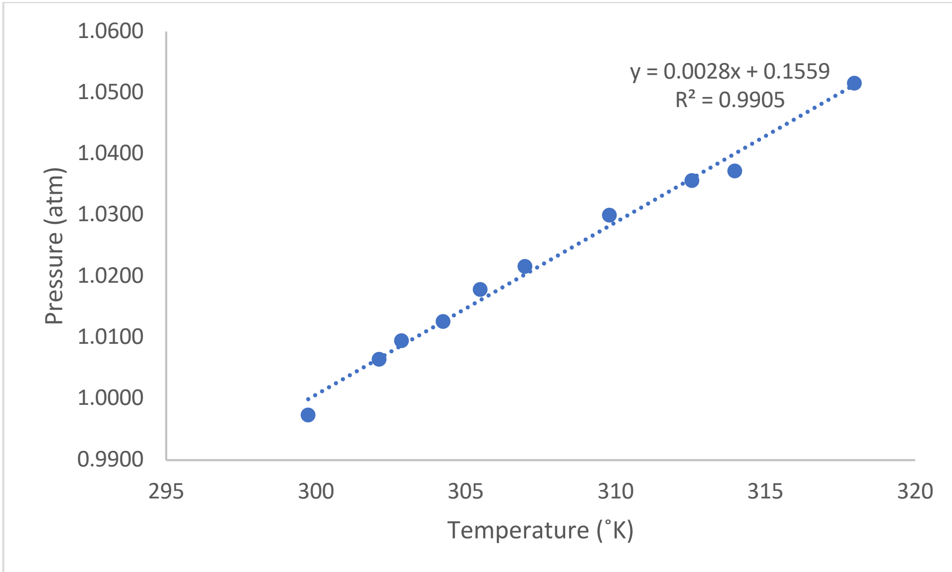



Figure 1: Relationship of Temperature with Pressure for air

| Part E: Relationahip of V vs n for a Particular Gas | | | | |
|---|---|---|---|---|
| Trial | H2O2 volume (mL) | Fe(NO3)3 volume (mL) | Theoritical moles of Oxygen gas produced | Volume of Oxygen gas (mL) |
| 1 | 0.68 | 5 | 0.01 | 5 |
| 2 | 1.36 | 5 | 0.02 | 25 |
| 3 | 2.04 | 5 | 0.03 | 40 |
| 4 | 2.72 | 5 | 0.04 | 50 |

Table 2: Experiment Data of Volume with Moles of Oxygen gas

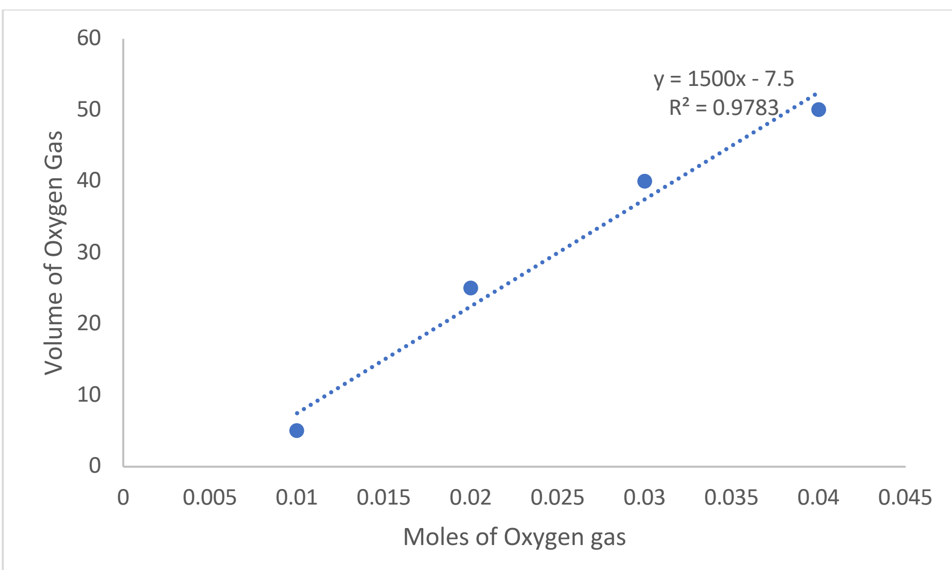



Figure 2: Relationship of Volume with Moles of Oxygen gas

*Conclusion*

Gas behavior on a macroscopic level can be described relations shared between its intrinsic properties. In particular, the behavior of gases can simply be determined by the current state of its pressure, volume, number of particles, and its temperature. In fact, the ideal gas law given below outlines the relationship between all 4 variables up to some constant, R.

$$PV = nRT$$

Specifically, during lab, we drew conclusions for the relationship of pressure with temperature for air. In our apparatus, we needed to measure one variable with respect to the other without introducing variance from other variables. We need to keep volume and the amount of air inside our container constant.

In the experiment, we decided to measure the pressure of the air with respect to its temperature. As such, the temperature is our independent variable, and the pressure is our dependent variable. When rearranging the ideal gas law to isolate P, we get:

$$P = \frac{nR}{V}T$$

Figure 1 illustrates a linear relationship between pressure and temperature and as such, the slope of the line of best fit is representative of the term:

$$\frac{nR}{V} = .0028$$

Additionally, in part E of our experiment, we measured the relationship of volume with respect to the number of moles of oxygen gas produced; the volume is dependent to the number of moles of oxygen gas produced. Applying this to the ideal gas law, we can rearrange the terms to isolate V and get:

$$V = \frac{TR}{P}n$$

Figure 2 shows us that there exists a linear relationship between the two variables for oxygen gas where:

$$\frac{TR}{P} = 1500$$

However, as you can see in figures 1 and 2, there is a nonzero y intercept term. In relation to the ideal gas law, this y intercept cannot be described and as such we can attribute this variance to the fact that these experiments were not carried out on ideal gases. The ideal gas law is applicable most accurately to gases in group 8 of the periodic table because they have a full shell of electrons and as such have zero valence electrons and thus the lowest reactivity. The air that we used for this experiment is comprised mostly of nitrogen, oxygen and carbon, all gases that are non-ideal. As such, much of the variance in this experiment is because the air is non-ideal.