GOAL
This experiment is designed to help the student better understand how the properties of gases relate to each other and how they can be used to find molar mass.

INTRODUCTION
In Part 1 of this experiment you will measure pressure, volume, and temperature of a gaseous sample. When combined with the value of R (0.082057 L atm/mole K) and the Ideal Gas Law, PV = nRT, you can then calculate the number of moles present. Of course, knowing the number of moles in a sample is a very powerful thing that allows us to find out any number of other things. In this experiment, our goal is to calculate molar mass. Since molar mass is simply the mass in grams of one mole, we need only divide the mass of the gaseous sample by the number of moles found from the Ideal Gas Law.

We will put an excess of a volatile liquid into a flask with a narrow opening and immerse the flask in a water bath above the boiling point of the liquid. As the liquid vaporizes and leaves the flask through a hole in the cap, it will sweep the air out along with it. At the point when all the liquid has just vaporized, we have a flask full of just gas phase unknown. The temperature of the gas is equal to the temperature of the water bath. The pressure of the gas is equal to atmospheric pressure because the flask is open to the atmosphere. The volume of the gas is equal to the volume of the flask because the excess sample left through the vented cap. This volume can be calculated from the mass of water that fills the flask and that water’s density. The mass of the gas is measured by subtracting the mass of the empty, dry flask from the final mass of flask and condensed vapor.

We are determining the molar mass of the liquid sample when it is a vapor (gas), open to the atmosphere, and at the boiling point of water. It is the point after all the liquid has vaporized but before the stopper is put back on. We cannot easily measure the temperature and pressure of the gas at this point, but we can easily measure the temperature of the water bath and atmospheric pressure. We will make the reasonable assumption that the gas is in equilibrium with its surroundings, which means that the temperature of the gas is the same as the temperature of the water bath, and the pressure of the gas is the same as atmospheric pressure. Our barometer will
measure air pressure in kilopascals. You will need to convert this to atmospheres using the conversion factor: \(1 \text{ atm} = 101.3 \text{ kPa}\).

In Part 2 of the experiment, you will observe the relationship between pressure and volume of a gas, with temperature and moles held constant. Finally, you will observe the relationship between pressure and temperature, with volume and moles held constant. This last investigation will allow you to estimate the value of absolute zero.

HAZARDS
The greatest hazard in this experiment is getting burned by the boiling water. Part 1 must be done in the hood because it releases unpleasant odors. Part 2 is done on the regular bench.

You will work individually this week. Your instructor will tell you whether you should start with Part 1 or Part 2.

PART 1: MOLAR VOLUME OF A VOLATILE LIQUID

Laboratory Observations and Data:
Be sure to make initial observations, such as color, odor, etc. on your unknown. Record what you do and what you observe throughout the experiment. Prepare a table with the following information:
- Unknown number
- Initial mass of flask and stopper
- Mass of flask, stopper, and condensed vapor
- Temperature of water bath, in °C
- Mass of flask, stopper, and water
- Barometric pressure measured in kPa (later you will convert this to atmospheres)

Procedure:
From the reagent bench obtain a special flask, rubber stopper, vented cap, and an unknown liquid. Weigh the empty, dry flask and stopper to the nearest milligram (0.001 g). Use a dropper to put your liquid (about 5 mL) in the flask.

The apparatus shown in Figure 1 is already assembled in the hood. Clamp onto the vented cap and use this to hold the flask below the liquid level. Heat to a gentle boil. Several boiling stones in the water will provide a more controlled boil. Careful observations will note a stream of vapor coming out the top and a decrease in the amount of liquid in the bottom.

When it appears that no more liquid is being boiled off, wait about 2 minutes and then measure the temperature of the boiling water and turn off the burner. After the water has stopped boiling, carefully unclamp the flask (do not remove it completely from the water) and immediately replace the vented cap with the rubber stopper used previously. Remove the flask from the beaker of water...
and immerse it in a beaker or pan of cool water. After about 2 minutes carefully remove the stopper for only one or two seconds to allow air to enter and again insert the stopper. Be sure you are recording what you do and observe in your lab notebook.

Gently dry the flask with a towel, loosen the stopper briefly to equalize any pressure difference, and re-weigh the flask. Read the atmospheric pressure from the barometer.

Fill the flask with distilled water, stopper it, and weigh it. You need the weight to +/- 0.01 g. Carefully remove the water. Return the flask and vented cap to the oven labeled “wet.” Return the beaker and rubber stopper to the hood.

PART 2: EXPLORING THE PROPERTIES OF GASES

Part 2 will be done using a Vernier computer setup on the open bench.

The relationship between pressure and volume
1. Position the piston of a plastic 20 mL syringe so that there will be a measured volume of air trapped in the barrel of the syringe. It is recommended to start with the syringe about half full of air. Attach the syringe to the valve of the Gas Pressure Sensor, as shown in Figure 2. A gentle half turn should connect the syringe to the sensor securely. Note: Read the volume at the front edge of the inside black ring on the piston of the syringe, as indicated by the arrow in Figure 2.

![Figure 2](image)

2. Connect the Gas Pressure Sensor to Channel 1 of the Vernier computer interface. Connect the interface to the computer using the proper cable.

3. Start the Logger Pro program on your computer. Open the file “Exp 6 Properties of Gases: Pressure and Volume.” This file allows you to collect pressure data from the Gas Pressure Sensor, using Events with Entry mode. For each pressure reading you take with a button, this mode lets you enter a volume value.

4. Measure the pressure of the air in the syringe at various volumes. The best results are achieved by collecting at least six data points. Do not disconnect the syringe between readings. Instead, change the position of the plunger to reach different volumes. When you have all your data, click the linear fit (R=) button at the top of the screen.

5. Print a copy of the graph and data.
The relationship between pressure and temperature

In this part of the experiment, you will study the relationship between the temperature of a gas sample and the pressure it exerts. Using the apparatus shown in Figure 3, you will place an Erlenmeyer flask containing an air sample in a water bath and you will vary the temperature of the water bath.

6. Connect the Gas Pressure Sensor to Channel 1 and a Temperature Probe to Channel 2 of the interface.

7. Assemble the apparatus shown in Figure 3. Be sure all fittings are airtight. Make sure the rubber stopper and flask neck are dry, then twist and push firmly on the rubber stopper to ensure a tight fit.

8. Open the file “Exp 6 Gases Pressure and Temperature.”

9. Take your 1 liter beaker to the sink. Allow the hot water to run until it is as hot as possible. Half fill your beaker with this hot water.

10. Open the valve on the top of your Erlenmeyer flask and immerse the flask up to its neck in the hot water. After holding it there for a minute or so, close the valve. This value must remain closed for the rest of the experiment. Double check to ensure that the tubing’s screw connections are all snug and that the valve on top is closed.

11. Click the Collect button.

12. Use one hand to keep the Erlenmeyer submerged. When both the temperature and pressure values seem to stabilize, collect a data point by clicking on the Keep button. [Both temperature and pressure will be dropping very slowly since the beaker is above room temperature, so the values will never to absolutely constant.]

13. Briefly remove your Erlenmeyer from the water bath and add a handful of ice to the warm water. Stir the water mixture with a glass rod until all the ice melts. Check the temperature. Continue adding ice until the water temperature is about 10 degrees cooler than your first measurement.

14. Again immerse the Erlenmeyer flask in the water up to its neck. Hold it there a minute or two. When both temperature and pressure values stabilize, collect another data point by clicking on the Keep button.

15. Continue this process of lowering the bath temperature and collecting new data points. Collect pressure data about every 10 degrees. Your final bath should be a slushy mixture of ice and water to reach 0°C. You may need to pour out some water along the way if the beaker becomes too full.

16. When you have all your data, click Stop, and then the linear fit (R=) button at the top of the screen. Print a copy of the graph and data. Be sure that you get the equation for your line.
FOR YOUR LAB REPORT

RESULTS and DISCUSSION
For Part 1, prepare a table with rows: unknown number, mass of the vapor in g, temperature in °C, pressure in kPa, volume of gas in mL, moles of gas, and molar mass. Note that the first four rows are simply data you collected in lab. The volume of the gas is equal to the volume of the flask. To find this, first calculate the mass of water that filled the flask. Then, assuming that the water’s density is 0.99802 g/mL, calculate the volume. The number of moles of gas is calculated using the Ideal Gas Law and your data from this table. Be careful; you will need to change the units on your data to match the units on R! Finally, the molar mass is just the mass of the gas divided by the number of moles. Show sample calculations for the calculation of volume, moles and molar mass. Directions for sample calculations can be found in the Introduction Handout on Inquire. Be careful to use correct units and significant figures.

Part 2 will be discussed through the Question section. Be sure to attach your data and graphs from Part 2 to your report.

QUESTIONS
1. Given below are three procedural errors that might be made in Part 1 of this experiment. State how each would affect the calculated molar mass by stating whether the molar mass calculated when the procedural error was made is greater than, equal to, or less than that calculated without the error. Explain each answer.

   a. On a day when the actual atmospheric pressure was 0.965 atm, a student simply assumed it would be about 1 atm and used that value in his calculations.

   b. When the flask plus condensed vapor was weighed, there was still some water on the outside of the flask.

   c. After all the sample had vaporized, but before the flask was removed from the water bath, air got into the flask and replaced some of the sample vapor. (Hint: The average molar mass of air is less than the molar mass of the unknown.)

2. H₂ gas is formed when HCl reacts with Mg metal.
   a. Write a balanced equation for the reaction of HCl with Mg.
   b. Calculate the mass of H₂ that will form when 0.23 g Mg is added to 25 mL of 1.0 M HCl. Don’t forget to check for limiting reagent. Show this calculation clearly in your report.
   c. What volume will the H₂ gas occupy at 21°C and 0.988 atm pressure? Show work.

3. Jim is upset. “I really screwed up Part 1 of this experiment. I just calculated the volume of my gas from the water data and got about 250 mL, but I know I only had two or three milliliters of sample, no where near 250 mL. My calculations must be all wrong.” Sara walks by and says, “You’re fine. Remember that you heated it up.” Who’s right? Can heating the sample make this
much difference? Choose a side. Explain what happens when the sample is heated, and why it happens, fully enough that you can convince Jim and Sara of the correct answer.

4. Consider your Part 2 graph of gas pressure and volume. Are these variables directly or inversely proportional? Explain how you can tell by looking at the graph. Which of the variables in $PV = nRT$ were held constant for this graph? Explain how this was accomplished in lab.

5. Consider your Part 2 graph of gas pressure and temperature. What is the equation for your line? Substitute into the equation to find temperature when pressure is zero. In theory, what should temperature be when pressure (or volume) is zero? How does this value affect the units you use for temperature when doing calculations with the Ideal Gas Law?