Experiment 11

Gas Laws

Pre-Lab Assignment

Before coming to lab:

- Read the lab thoroughly.
- Answer the pre-lab questions that appear at the end of this lab exercise.

Purpose

Boyle's Law and Gay-Lussac's Law will be experimentally confirmed by preparing graphs and observing the relationship between the data. The molar mass of acetone will also be experimentally determined using the Ideal Gas Law.

Background

In the 17th century, physicist Robert Boyle first published what became as Boyle's Law, stating that for a fixed amount of gas at a constant temperature, the product of pressure and volume is constant. This means that the pressure and volume of a gas are inversely proportional. This can be restated to calculate changes to the pressure or volume of a gas, as seen in Eqn. 1.

$$PV = \text{constant or } P_1V_1 = P_2V_2 \qquad \text{Eqn. 1}$$

Example Problem: Using Boyle's Law

A sample of gas is held at constant amount and temperature at 561 mmHg occupies 0.50 L. Calculate the new volume, in L, if the pressure is increased to 0.98 atm.

Step 1: Convert pressure into the same unit.

$$0.98 \text{ atm} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} = 740 \text{ mmHg}$$

Step 2: Use Eqn. 1 to find V_2.

$$(561 \text{ mmHg})(0.50 \text{ L}) = (740 \text{ mmHg})(V_2)$$

$$\frac{(561 \text{ mmHg})(0.50 \text{ L})}{740 \text{ mmHg}} = V_2 = 0.38 \text{ L}$$

In the early 1800s, Joseph Louis Gay-Lussac stated that when the amount and volume of a gas were held constant, the pressure and temperature were directly proportional. This law can be stated mathematically two ways, as seen in Eqn. 2.

$$\frac{P}{T} = \text{constant or } \frac{P_1}{T_1} = \frac{P_2}{T_2} \text{ or } P_1T_2 = P_2T_1 \qquad \text{Eqn. 2}$$

Kinetic Molecular Theory states that the average energy of molecules in the gas phase is directly proportional to the gas's absolute temperature. The pressure a gas exerts is due to the collisions

Example Problem: Using Gay-Lussac's Law

A sample of gas held at a constant amount and volume exerts 625 torr at 31.0°C. Calculate the temperature, in K, when the pressure reduces to 595 torr.

Step 1: Convert all temperatures to Kelvin.

31.0°C + 273.15 = 304.2 K

Step 2: Use Eqn. 2 to find T_2.

$$\frac{625\text{ torr}}{304.2\text{ K}} = \frac{595\text{ torr}}{T_2}$$

$$T_2 = \frac{(595\text{ torr})(304.2\text{ K})}{625\text{ torr}} = 2.90\times10^2\text{ K}$$

of its molecules with the sides of the container. Therefore, if a gas's kinetic energy approaches zero, its temperature and pressure must also approach zero.

The process of identifying unknown compounds is a cornerstone of empirical laboratory procedure. One of the most important pieces of information needed is the compound's molecular weight. With this in hand, the chemist can determine the molecular formula of the compound and therefore also determine the structure and begin predicting physical and chemical properties it may exhibit.

There are a variety of techniques that can help to determine a compound's molecular weight. This experiment will utilize the Dumas Bulb Method—a very simple experiment that involves the Ideal Gas Law (PV = nRT) and Dalton's Law of Partial Pressure ($P_T = P_1 + P_2 + P_3 ... + P_n$). A volatile liquid is placed into a closed container and then heated until all the liquid has vaporized and a mixture of air and gas fills the container. The total pressure, temperature, and volume of both the air and the now-volatilized liquid are measured. These values are used to calculate the amount of vapor in moles present in the flask, given its partial pressure, volume, and temperature. Since the original mass of the sample was known, the molar mass can be calculated.

Example Problem: Finding the molecular weight of an unknown

An empty Erlenmeyer flask was heated to 68.5°C and had a pressure of 100.8 kPa. When 0.10 mL of methanol (density = 0.792 g/mL) is added to the flask and fully vaporized, the pressure increased to 132.5 kPa. The total volume of the flask was 132.4 mL. Calculate the molecular weight of methanol.

Step 1: Find the pressure of methanol vapor

132.5 kPa – 100.8 kPa = 31.7 kPa

Step 2: Convert all units to use in the Ideal Gas Law

$$31.7 \text{ kPa} \times \frac{1 \text{ atm}}{101.325 \text{ kPa}} = 0.313 \text{ atm}$$

$$132.4 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.1324 \text{ L}$$

68.5°C + 273.15 = 341.7 K

Step 3: Solve for moles in the Ideal Gas Law

(0.313 atm)(0.1324 L) = n (0.08206 L*atm/mol*K)(341.7 K)

$$n = \frac{(0.313 \text{ atm})(0.1324 \text{ L})}{(0.08206 \text{ L*atm/mol*K})(341.7 \text{ K})} = 0.00148 \text{ mols}$$

Step 3: Find the mass of the unknown

$$0.10 \text{ mL} \times \frac{0.792 \text{ g}}{1 \text{ mL}} = 0.079 \text{ g}$$

Step 4: Find the molar mass

$$\frac{0.0.79 \text{ g}}{0.00148 \text{ mols}} = 53 \text{ g/mol}$$

Procedure

Part I: Boyle's Law

1. Obtain a pressure sensor kit.

2. Connect the pressure sensor to the LabQuest 2 or computer. If using a computer, start Graphical Analysis before connecting the sensor. Do not connect the syringe yet. When connected, the screen should automatically show the current pressure from the sensor in kPa.

3. Tap Sensors, then Data Collection. Change the data collection mode to "Events with Entry" from the drop-down menu. In the Name box, enter "Volume" and in units, "mL". Click the Graph icon in the top right-hand corner. The graph should be updated to Volume (mL) versus Pressure (kPa) (Fig. 2).

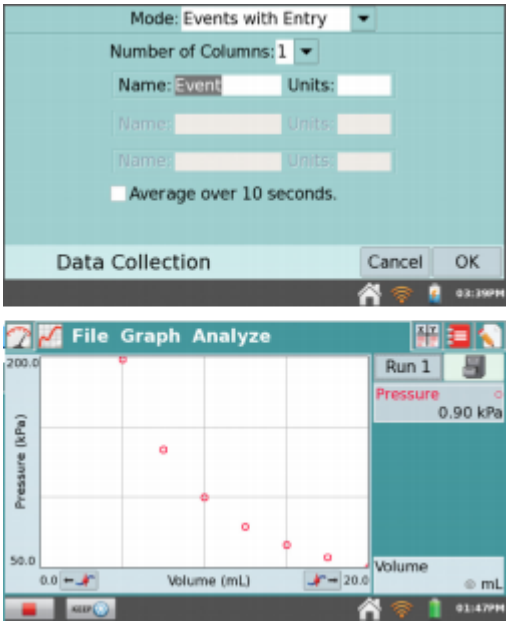

Fig. 2: Settings and Graph for Boyle's Law

4. Using the plastic syringe included in the pressure sensor kit, move the piston until the syringe reads 20.0 mL. Hold it in place while screwing the syringe onto the fitting on the pressure sensor (Fig. 3).

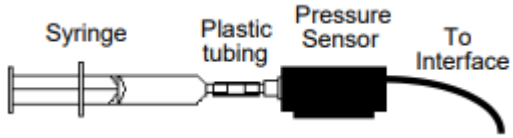

Fig. 3: Pressure sensor set-up

5. Press the Green Triangle (Start) in the bottom left-hand corner. The sensor will not record any data until the Keep button is pressed.

6. Press Keep. A pop-up window will appear. Enter the volume from the syringe.

7. Carefully move the syringe's piston up to the 18 mL mark. Hold it in place and press Keep. Enter the new volume of the syringe in the pop up window.

8. Repeat Step 7 for 2 mL intervals (16, 14, 12, etc.) until at least 10.0 mL or further if you are able. When it becomes too difficult to depress the syringe further, press the Red Square (Stop) in the bottom left-hand corner..

9. Print the graph. If using a LabQuest 2, export the graph and data via a flash drive to a computer and print it.

10. Prepare a second graph of Inverse Volume (1/V, x-axis) versus Pressure (y-axis). Print this graph.

Part II: Gay-Lussac's Law

1. Obtain a pressure sensor kit and a glass thermometer.

2. Connect the pressure sensor to the LabQuest 2 or lab computer. If using a computer, make sure Graphical Analysis is open before connecting the sensor. Connect the tubing with rubber stopper to the pressure sensor. You should see the current pressure reading on the screen.

3. Tap Sensors, then Data Collection. Change the data collection mode to "Events with Entry" from the drop-down menu. In the Name box, enter "Temperature" and in units, "°C". The graph should update to Temperature versus Pressure.

4. Fill a 600 mL beaker approximately half-full with a mixture of tap water and ice. Set up a Bunsen burner, wire gauze, and iron rings. Put the beaker on top of the wire gauze above the Bunsen burner.

5. Gather a dry 125 mL Erlenmeyer flask. Put the rubber stopper connected to the pressure sensor tubing firmly in its top. Make sure that the valve on the short tubing is closed. Gently lower this into the water bath from Step 4. You will need to clamp the neck in place to submerge the flask as far as possible without allowing any water to seep over the edges into the flask itself.

6. Press Start on Graphical Analysis. The sensor will not record any data until the Keep button is pressed.

7. With a split stopper slid onto the thermometer, clamp the glass thermometer inside the water bath about 1 cm from the bottom. Do not allow the bulb to touch the edges of the beaker or Erlenmeyer flask. Allow the system about one minute to come to thermal equilibrium before pressing Keep. A pop-up window will appear. Enter the temperature listed on the thermometer.

8. Turn on the Bunsen burner and heat the water until the temperature increases by approximately 20°C. Turn off the burner and press Keep. Enter the temperature listed on the thermometer.

9. Repeat Step 8 three more times at approximately 40, 60, and 80°C. **Do not allow the water to boil**. You should have a total of five total data pairs.

10. Press Stop.

11. Transfer your data to Microsoft Excel.

12. Convert each temperature to Kelvin.

13. Prepare a graph of Temperature (K, x-axis) versus Pressure (y-axis). Add a linear trendline and report the equation.

Part III: Molecular Weight of Acetone

1. Use the same set-up as in Part II, including the pressure sensor, interface, dry 125 mL Erlenmeyer flask, tubing and stopper, and 600 mL beaker with water bath set up over a Bunsen burner with thermometer in place.

2. Tap Sensors, then Data Collection. Change the data collection mode to "Time-Based". Set the Rate to 10 samples/second and the Duration to 20 minutes.

3. Press the Green Triangle (Start) button in the bottom left-hand corner. Click on the Graph icon to watch the data being collected.

4. Heat the water bath to approximately 60°C. The exact temperature is unimportant so long as it is always above 56.5°C (the boiling point of acetone).

5. Measure 0.20 mL of acetone with a microsyringe. Place the microsyringe in the small rubber tubing with the valve still closed.

6. When the water bath reaches 60°C, allow 30 seconds for it to achieve equilibrium (the pressure and temperature should be constant) before opening the valve and carefully squirting all of the acetone into the Erlenmeyer flask. Close the valve.

7. Continue collecting data until the system reaches its new equilibrium. Click the Red Square (Stop) button in the bottom left-hand corner.

8. Record the temperature on your data sheet.

9. Remove the stopper. Fill the Erlenmeyer flask to the brim with tap water and measure the volume of the water in the flask using your graduated cylinder

Note: You will need to fill your graduated cylinder, record the volume, and then empty it repeated times and then add each volume to find the total. It will not be 125 mL.

10. Use the Ideal Gas Law to solve for moles of acetone.

11. Look up the density of acetone and use it to calculate the grams of acetone used.

12. Calculate the molecular weight of acetone.

14. Compare your calculated molecular weight of acetone to the tabulated value and report its percent error.

Experiment 11—Data Sheet

Name: _____

Part I: Boyle's Law

Attach both graphs of Volume versus Pressure and 1/Volume versus Pressure.

Part II: Gay-Lussac's Law

Temperature (°C)	Temperature (K)	Pressure (kPa)

Attach the graph of Temperature (K) versus Pressure.

Trendline Equation: _____

Part III: Molar Mass of Acetone

1. Pressure of air alone before injection (kPa) _____

2. Pressure of air and acetone after injection (kPa) _____

3. Temperature after injection (°C) _____

4. Volume of flask (mL) _____

5. Change in pressure after injection (kPa) _____
show calculation:

6. Change in pressure after injection (atm) _____
show calculation:

8. Temperature after injection (K) _____
show calculation:

9. Volume of flask (L) _____
show calculation:

10. Moles of acetone (mol) _____
show calculation:

11. Grams of acetone (g) _____
show calculation:

11. Molar Mass of acetone (g/mol) _____
show calculation:

12. Tabulated Molar Mass of acetone (g/mol) _____

13. Percent Error (%) _____
show calculation:

Experiment 11—Post-Lab Assignment

1. From your graphs from Part I, is the relationship between pressure and volume of a gas *directly* or *inversely* proportional? Does this agree with Boyle's Law? Explain.

2. Explain why one of your graphs from Part I is linear while the other is curved.

3. From Part I, what is the slope of the line in your graph of 1/Volume versus Pressure? (Hint: this is not a numerical value. Define it in terms of the variables of the Ideal Gas Law, PV = nRT).

4. Give three experimental errors or limitations in Part I and decide what effect (too high, too low, no effect) they would have on the recorded pressure and volume values.

5. A syringe at STP has a total volume of 25.0 mL. If the syringe is depressed to a volume of 18.5 mL, what will the new pressure be inside the syringe, in atm?

6. From your graph in Part II, is the relationship between pressure and volume of a gas *directly* or *inversely* proportional? Does this agree with Gay-Lussac's Law? Explain.

7. Using the trendline equation from your graph in Part II, calculate at what temperature, in Kelvin, the pressure inside the flask will be 0 kPa.

8. What is the theoretical temperature at which pressure inside the flask should equal 0 kPa (Question 7)? Calculate your percent error.

9. Give three experimental errors or limitations in Part II and decide what effect (too high, too low, no effect) they would have on the recorded pressure and temperature values.

10. If a gas held at 1.2 atm at 25°C is heated to 52°C, what will the new pressure of the gas be, in atm?

11. Is Dalton's Law of Partial Pressure used in this experiment? If so, explain how and where you used it.

12. From Part III, was your percent error in the molar mass of acetone large or small? Give three possible experimental causes for this difference.

Experiment 11—Pre-Lab Assignment

Name: _____

For all calculations, show all work and draw a box around the final answers.

1. What is the relationship between the pressure and volume of a gas?

2. What is the relationship between the pressure and temperature of a gas?

3. A 2.50 L flask containing $N_2(g)$ at a pressure of 0.63 atm is directly connected to an empty 1.25 L flask with the valve initially closed. When the valve is opened, the gas is allowed to equilibrate freely between both flasks. Calculate the new pressure of the gas, in atm.

4. A sample of He(g) is kept in a container with a fixed volume at 23.8°C and at 0.971 atm. If the pressure of the gas is decreased to 0.828 atm, will its temperature increase or decrease? Calculate the new temperature of the gas at this pressure.

5. What does the term "volatile" mean? Why is it necessary for acetone to be volatile for this experiment?