Experiment VII: Ideal Gas Laws

Goals

- Study the relationship between pressure and temperature of atmospheric gas at constant volume; Determine the absolute zero temperature in Celsius

- Study the relationship between pressure and volume at constant temperature

Introduction and Background

Early studies of gases revealed that in most cases the relationship between pressure, volume, and temperature was essentially independent of the kind of gas involved, as long as the conditions were not close to leading to condensation. Significant advances were made by Boyle, Charles, and Gay-Lussac.

Charles discovered that when volume was plotted versus temperature at constant pressure, a straight line resulted, namely, $V = C \times T$. When plots for different gases were compared, they all extrapolated to give zero volume at the same absolute temperature. This is referred to as Charles' Law.

Gay-Lussac studied the relation between pressure and temperature under constant volume and found a linear relationship between the two: $P = C' \times T$. Again the extrapolation to zero pressure gives the same temperature for different gases. This is now called the Gay-Lussac's Law.

Boyle investigated the relation between pressure and volume at constant temperature and found that $PV = const$. This is the Boyle's Law.

Combining the results of the three investigators, a universal ideal gas law was derived: $PV = nRT$, where n is the number of moles and R the ideal gas constant.

In this lab we will verify the Gay-Lassac's law and Boyle law for simple atmospheric gas, and we will use the data for the Gay-Lussac's experiment to determine the absolute zero temperature on the Celsius scale.

Experimental Setup

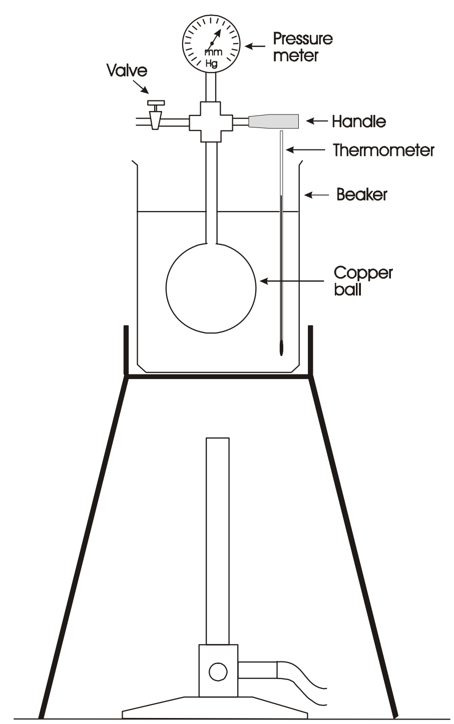

Figure 7.1 – Pressure vs. temperature apparatus.

Equipment**:** Copper ball with valve and pressure gauge, 10 cm^3 syringe, Bunsen burner, Stainless steel beaker, three-leg support, thermometer, weights, table clamps, caliper.

Setup**:** The experimental setup for the Gay-Lussac's experiment is shown schematically in Figure 7-1. A copper ball containing a fixed amount and volume of atmospheric gas is connected to a pressure gauge. It is immersed in a water bath so that its temperature can be varied between 0^oC and 100^oC.

For the Boyle's experiment, a syringe with a shut-off valve is used. A fixed amount of gas is trapped in the syringe and its temperature is maintained at room temperature. The pressure of the gas is changed by hanging different weights on the cylinder.

Experimental Procedure

A. Gay-Lussac's Law

1. Open and then close the valve on the copper ball at room temperature. Read the pressure and the room temperature. Record these readings in a two-column table for pressure and temperature.

2. Immerse the copper ball in an ice-water bath with enough ice to reach an equilibrium temperature between 0 – 10^oC. *Make sure that the copper ball does not touch the bottom of the beaker and the ice-water covers the entire ball*. Place the thermometer in the bath and make sure that it does not touch the container. After the temperature has stabilized, record the temperature and the pressure into the table.

   Now heat the bath with a Bunsen burner. Record the pressure when the pressure gauge needle is at a tick mark. Tap the glass on the pressure gauge periodically, especially before taking a reading. At the same time record the temperature. Repeat until the water is boiling.

B. Boyle's Law:

1. Open the valve at the end of the syringe and set the plunger so that there is 9 cm^3 of air inside. Close the valve carefully. Record the volume and pressure in the table as shown below. Since what you are interested in is net pressure, you should record the initial atmospheric pressure as zero.

Mass (kg)	Force (N)	Pressure (N/m^2)	Volume (m^3)	Inverse V (m^{-3})
0	0	0	9×10^{-6}	1.11×10^{5}
0.200				
0.400				
…				

2. Apply a 200 g mass to the plunger. Read the volume of the air from the syringe. Record the mass and volume in the table above. Increase the mass by a 200-g increment and record the corresponding volume. Repeat the process until the mass is about 1.6 kg. The pressure in

the syringe can be calculated from dividing the applied force by the area of the plunger. Measure the diameter of the plunger with a caliper and calculate it area from it.

Data Analysis

A. Gay-Lussac's Law

Open the template named "Gaslaw" in Excel. The template has preset *x*-axis going from 0 to about 1000 mmHg and the *y*-axis going from -300^{o}C to 100^{o}C. From the slope and *y*-intercept of the linear regression fit calculate the temperature at which the pressure is zero. This is the absolute zero temperature.

- What is your value in ^{o}C for the absolute zero? Is your result consistent with the accepted value?

- What do you see as the major source of systematic error in this experiment?

  Gay-Lussac's law applies to pressure versus temperature at a fixed volume. Is it reasonable to neglect the volume expansion in the copper ball as the temperature is increased from 0 to 100 ^{o}C? (Use the following to answer the question: The volume expansion coefficient, $\beta = \frac{\Delta V}{V}\Delta T$, of copper is about 50 x 10^{-6} /^{o}C.)

  An additional source of error is the volume of air in the tube connecting the copper ball to the pressure gauge. Estimate the volume of air in the tube. When the temperature of the air in the copper ball is increased, the temperature of the air in the connecting tube may not increase by the same amount. How would this affect the pressure reading?

***B. Boyle's Law*:**

1. Use the second part of the "Gaslaw" template to plot pressure versus inverse volume on a linear scale. Find the slope of this plot from a linear regression.
   The ideal gas law gives PV = nRT, thus the slope of the plot gives nRT. Use the slope you have obtained to estimate the number of gas molecules contained in the syringe? (R = 8.315 J/mol*K and the Avogadro's number $N_A = 6.02 \times 10^{23}$ mol^{-1})

2. The number of molecules in the syringe can also be calculated directly by using the fact that at standard temperature and pressure (STP) the volume occupied by 1 mol of gas is 22.4 liters. What is the number of liters occupied by 1 mol of gas at room temperature under atmospheric pressure? From that estimate the number of molecules contained in the 9 cm^3 of air. Does your result agree with that you obtained from the slope?

3. What do you think are the major sources of error in this experiment? How does the friction between the plunger and the syringe affect the pressure? If there is a small leak in the valve, how would it affect your result?

Conclusions

Briefly discuss whether you have accomplished the goals listed at the beginning.