# Experiment 9A

FV 8/4/04

# EXPLAINING A SIMPLE GAS EXPERIMENT

**MATERIALS**: 125-mL Erlenmeyer flask, rubber stopper fitted for thermometer and with one additional hole, buret, tall beaker, ice, ruler

**PURPOSE**: The purpose of this experiment is to demonstrate aspects of the ideal gas law, measure pressures based on heights of liquid columns, and correctly account for the vapor pressure of water.

**LEARNING OBJECTIVES:** By the end of this experiment, the student should be able to demonstrate the following proficiencies:

1. Use a barometer to measure the ambient pressure.
2. Interconvert pressure units, including units relating to heights of liquid columns
3. Properly include the vapor pressure of water in calculations
4. Develop a calibration curve using experimental data and a spreadsheet
5. Demonstrate gas law relationships, including Boyle's Law and Charles' Law

**DISCUSSION**:

*Barometers and manometers*. Barometers and manometers function based, in part, on the principle of hydrostatic pressure. A column of liquid, of height h, exerts a pressure at its base, due to gravity, equivalent to $hg\rho$ , where g is the constant of gravitational acceleration (980.7 $cm/sec^2$) and $\rho$ is the density of the liquid. Of course, there may be additional sources of pressure above the column of liquid, e.g., the atmosphere, and these must be added to the pressure contribution of the liquid itself. Pressure units represent equivalencies of methods of describing or measuring pressure, e.g., 1 atm = 760 mm Hg = 14.7 $lbs/in^2$ .

*Vapor pressure*. Above every liquid is a vapor which is produced as molecules in the liquid state escape to the gaseous state. If the liquid is in a *closed* container, as is often the case in chemistry applications, the vapor will attain a pressure known as the *equilibrium vapor pressure*. Different liquids display unique equilibrium vapor pressures, depending on the magnitude of the intermolecular forces between the liquid state molecules. In addition, a given liquid's equilibrium vapor pressure increases with temperature. For example, the equilibrium vapor pressure of water is about 20 mmHg at room temperature, but 760 mmHg at $100^oC$. Tables or theoretical equations (e.g., the Clausius Clapeyron equation) can be used to determine vapor pressures at various temperatures.

**PROCEDURE**:

1. Observe the use of a barometer, as demonstrated by your instructor, including steps required to yield correct pressures. Take any important notes for later use.

2. Follow the procedures outlined below:

   a. Measure the barometric pressure.

   b. Add a few milliliters of distilled water to the Erlenmeyer flask. Weigh the flask. See Figure 1.

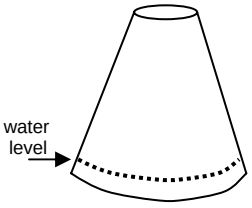


Figure 1

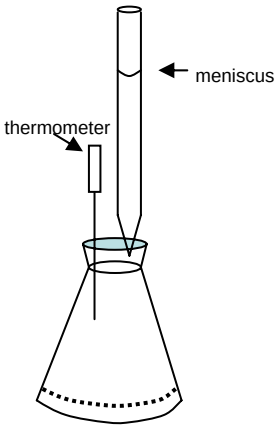


Figure 2

   c. Insert the rubber stopper with thermometer in the mouth of the flask snugly. Fill the buret with distilled water, removing any air from the buret tip. Carefully place the buret tip in the hole of the rubber stopper, *just barely snugly*. Make sure the buret valve is closed. Record the level of the water meniscus in the buret. See Figure 2.

   d. Open the valve of the buret and allow the water to freely flow into the flask until it stops (i.e., until at least 30 seconds between drips). Close the valve. Record the level of the water meniscus in the buret. Record the temperature. See Figure 3.

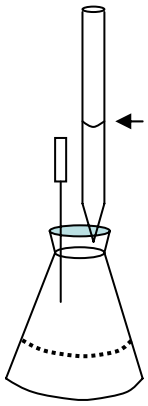

Figure 3

e. Immerse the flask in an ice bath, right up to the top half-inch of the rubber stopper. Wait until the temperature stabilizes. Record the temperature.

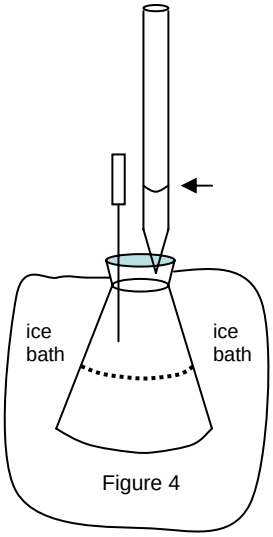

Figure 4

f. Open the valve on the buret and allow the water to freely flow into the flask until it stops, as before. Close the valve and record the meniscus level. See Figure 4.

g. Remove the flask from the ice bath and dry off the outside of the flask as much as possible. Carefully remove the buret from the rubber stopper.

h. Remove the rubber stopper and fill the flask completely with water. Insert the rubber stopper exactly as before, displacing any excess water. Then insert the buret exactly as before, displacing any excess water. Dry off the flask and stopper. Remove the buret and stopper carefully, making sure that all of the water drops stay in the flask. Weigh the flask plus water. See Figure 5.

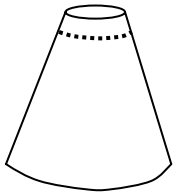
Figure 5

i. Based on experimental ruler measurements, use a spreadsheet to determine a calibration curve which relates the buret volume readings with the height above the tip of the buret.

**QUESTIONS FOR CONSIDERATION:**

1. Using your data, the calibration curve relating buret volume readings with water height above the buret tip, a table showing the density of water at various temperatures, a table of vapor pressures of water at various temperatures, and ideal gas law relationships, determine the following:

   a. the pressure of the gases inside the flask at the end of step c.

b. the pressure of the gases inside the flask at the end of step d, as well as the pressure at the buret tip due to the atmosphere and the water in the buret. Should these calculated pressures agree? Do they? Explain.

c. the pressure of the gases inside the flask at the end of step f, as well as the pressure at the buret tip due to the atmosphere and the water in the buret. Should these calculated pressures agree? Do they? Explain.

2. Suppose that a similar experiment is carried out using mercury liquid instead of water in both the flask and the buret. All other things being equal, would you expect the mercury level in the buret to fall more or less than the water level before equilibrating? Explain.

3. Why is it important to have some water in the flask in step b?