Experiment 12
Determination of Molar Mass Using the Ideal Gas Law

One of the most important applications of the Ideal Gas Law is the experimental determination of the molar masses (or "molecular weights") of gases. In fact, the method used in this experiment was the first widely-used method for determining molar masses, and was extremely important in the development of chemical theory.

The molar mass is the mass of one mole of a substance, so to determine this value, we must know two things: how many moles we have, and how many grams the sample weighs. The first thing - number of moles - can be calculated for a gas if a known volume of gas is present at a known temperature and pressure. This is obtained by vaporizing a sample in a flask of measured volume. The second piece of information - the mass of the sample - is obtained by weighing the condensed vapor.

The "unknown" substances provided for this experiment are liquids at room temperature. However, they are easily converted to gases (vapors), since their boiling points are less than 100 °C. This is accomplished when a small amount of liquid sample is placed in the flask, which is then heated in a water bath. The liquid sample vaporizes, and we assume that the temperature of the vapor formed is the same as the temperature of the warm water bath that surrounds the vapor. The pressure of the vapor in the flask must equal the pressure of the atmosphere at the time, since the flask is not closed off against atmospheric air. Enough of the liquid sample is used so that its vapor will be enough to more than fill the sample flask. As a result, the vaporizing sample forces the air out of the flask as the vapor itself fills the flask. Any excess vapor also leaves the flask. Thus, after a few minutes, the flask contains only the vapor, so we know that the vapor's volume equals the volume of the flask. At this moment, we record the water bath's temperature, which we take to equal the sample's temperature.

Next, the flask is removed from the warm water. Of course, the sample cools, and the vapor condenses back into a liquid inside the flask. The fact that it is now a liquid does not matter; its weight will still be the same. Subtraction of the weight of the flask itself from the weight of the flask with sample in it will give the sample weight.

Safety Precautions:

- Wear your safety goggles.

Waste Disposal:

- Discard any excess unknown liquids in one of the **organic waste** containers (with a pink label) in one of the fume hoods.

Procedure

1. Loosely cover a clean, dry 250-mL Florence flask with either a square of aluminum foil or a small square of clinging plastic food wrap. Weigh the flask with the cover on it. Do not try to measure the volume of the flask at this time.
2. Remove the cover from the flask and add a small sample of liquid unknown - about 3 mL - to the flask. (Knowing the exact volume is not important - why?) Replace the cover and crimp it tightly around the neck of the flask. Avoid leaving any loose foil or plastic wrap hanging down that might collect moisture from the water bath during the experiment.
3. Use a pin to make a tiny pinhole in the center of the foil or plastic wrap cover. This is done to allow the air and excess vapor to leave the flask, maintaining a pressure equal to atmospheric pressure.

4. Set up a ring stand with a ring and a piece of wire screen over a Bunsen burner. Adjust the height of the ring so that the hottest part of the flame will heat the beaker that is on the ring stand. (If the beaker of water is too far away from the flame, it will take a very long time for the water to heat up!) Clamp the flask at its very top, and suspend it in a 1- liter beaker of tap water on the ring stand. Without allowing any water to get into the flask, push the flask as deep into the bath as possible, since the entire vaporized sample should be surrounded by water. (Why? To make sure that the entire sample is the same temperature throughout.)
5. Begin heating the bath with a burner, and occasionally stir the water with the thermometer. Be very careful with the thermometer! Leave the thermometer in the bath throughout the warming time.
6. Continue warming, and watch the sample continuously. It may be necessary to loosen the clamp at the end attached to the ring stand and gently shake or swirl the flask <u>while keeping it submerged</u> to enable you to see the sample. The sample may boil, but it might instead vaporize entirely without ever getting warm enough to boil.
   It is very important that you keep the sample submerged during the vaporization process - do not take the flask out and put it back in the bath before the vaporization is complete, or it may jeopardize the accuracy of your results.
7. When you are certain that the sample has vaporized entirely, record the temperature of the bath and remove the sample flask from the bath.
8. Let the flask hang on the ring stand until it has cooled to room temperature. During the wait, read the barometric pressure and record it.
9. When the flask has cooled, wipe off all traces of water that is on the outside of the flask, and then weigh the flask (and sample) with the cover still in place.
10. For your second trial, add about 2-3 more mL of the same unknown to the flask (Do not dump out the sample remaining at the end of step 9! Why?) Replace the cover (make sure it is the same cover!), and **repeat steps 4-9.** Use the same water bath as before - since the water is already warm, it will take less time to heat up.
11. You have now recorded temperature, pressure, and weight of the sample for two separate trials. Volume must be measured next. This is NOT the same as the "250-mL" marked on the flask! To determine the volume, remove the cover and dump out the sample into the organic waste. Then, rinse out the flask and then fill it to the brim with tap water. Pour this water into a 500-mL graduated cylinder, and measure the volume to the nearest milliliter.

<u>Calculations</u>

As outlined in the introduction, calculate the number of moles of sample that were present for each trial using the Ideal Gas Law (PV = nRT). Assume that the pressure of the gas at the time you take the temperature measurement is the same as the atmospheric pressure, and that the volume of the gas is equal to the volume of the flask. Then, using that number of moles and the weight of the sample, calculate the molar mass for each trial. Be sure that your units are consistent with the value of R (the ideal gas constant) that you use.

Calculate a separate value of molar mass for each trial. Calculate the average value and the percent difference between the two values. If you trust one result more than the other, state which one and explain your reasoning.

Questions

1. Each of the following errors or facts would cause the molar weight that you calculate to be either larger or smaller than the true value. For each, you are to explain whether the calculated (wrong) value would be larger or smaller than the true value, and WHY.
   a. Suppose that you took the flask out of the bath before the entire sample had vaporized.
   b. Suppose that the sample vapor's temperature lagged behind the temperature of the bath, so that you were calculating with a temperature figure higher than the true one.
   c. Suppose you forgot to dry off all the water before weighing the flask.
   d. Suppose that you took the flask out of the bath before the entire sample had vaporized, and then put it back in the bath to finish vaporizing.

2. To calculate how many moles of vapor were present in the flask, you used the Ideal Gas Law, but in fact, the vapor is probably not acting like an Ideal Gas under the conditions of the experiment. Why not? At this temperature, not far from the boiling point of the liquid, are the molecules likely to be closer together than ideal, or further apart than ideal? How would this affect the mass of unknown gas that would fit into the flask at this temperature and pressure? How would it affect the value of molar mass obtained? (Again, explain whether the calculated value would be larger or smaller than the true value, and why.)

3. At the end of the experiment, you weigh your unknown in the flask after it has condensed back into a liquid.
   a. Why don't you just weigh the unknown as a gas?
   b. How do we take into account the air in the flask?
   c. Why do you put in excess unknown liquid before the experiment, and why don't you weigh it or measure its volume before heating it?

4. In your laboratory report, describe the assumptions that we are making for the calculations of this experiment.