

Lab 4: Gas Laws

Determining the Molar Mass of a Low-Boiling Liquid

Introduction

In this experiment, you will determine the molar mass and identity of a liquid that boils at a low temperature. The method you will use was developed by JBA Dumas (1800-1884), a French chemist who was an early contributor to the theory of molecules. The Dumas method involves heating a liquid until it changes to the gaseous state, at which point you can measure the pressure, volume, and temperature of the gas. These measurements, along with the mass of the gas can be combined in the ideal gas law to calculate the molar mass of the sample.

The ideal gas law, derived in section 5.3 of the Zumdahl text, is an empirically determined equation. This means that many experiments were performed in order to observe and better understand the behavior and properties of gases. When the gas laws of Boyle, Charles, and Avogadro (Zumdahl, section 5.2), relating the volume of a gas to pressure (P), temperature (T), and the number of moles of gas (n), respectively, are combined with the universal (ideal) gas constant, the resulting expression is the ideal gas law:

$$PV = nRT$$

where R is $0.08206 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$. It is important to note the units of the gas constant because any values plugged into the equation must use the same units (Pressure-atm, Volume-L, Temperature-K).

The ability to algebraically manipulate the ideal gas law makes it very useful and easy to apply in the lab. For instance, the number of moles (n) of a substance can also be written as

$$\frac{\text{mass}}{MM} \quad \text{since} \quad \frac{\text{g}}{\text{g/mol}} = \text{mol}$$

Rewriting $PV = nRT$ using mass gives you:

$$PV = \frac{m}{MM} RT$$

which can also be rearranged in order to determine the density (mass/volume) of the gas of interest:

$$\frac{m}{V} = \frac{P * MM}{RT}$$

Most gases follow the ideal gas law at low pressures (< 1 atm) and high temperatures. We will use this law to determine the molar mass and identity of an unknown liquid by heating a sample of the liquid and recording the pressure, volume, and temperature of the sample in the gaseous state, as well as its mass. After plugging this information into the ideal gas law and

calculating the molar mass of the sample, you will identify your unknown from a list of possible compounds.

Helpful information

Once you have vaporized your sample, you will remove the heat source and immediately cap the flask to trap the vapor inside. The pressure and temperature at that moment, along with the volume and mass of the sample will give you the information necessary to calculate the molar mass of your compound using $PV = nRT$. If you know P , V , R and T , you can solve for n (number of moles). Then, dividing the mass of the sample by the number of moles will give you units of g/mol, which is the molar mass.

That said, the pressure and temperature of your sample are easily obtained from the barometer on the wall (instructions are provided in Appendix D of the lab manual) and the temperature of the water used to vaporize your sample (boiling - 100°C). The gas constant, R , is

$$0.08206 \frac{L \cdot atm}{mol \cdot K}$$

The volume and mass of your vaporized sample take a little bit more work to figure out.

Volume. The volume of the vapor is the volume of flask in which you trap it. Since each flask in the lab is not exactly the same volume, each group will need to determine the volume of the flask they use. To do this, you will fill the flask with water, determine the mass of the water (flask/cap/water – flask/cap) and, knowing the temperature of the water, use the density of water at that temperature to determine the volume of the flask.

$$vol. of the flask = (g \text{ water}) \left(\frac{mL \text{ water}}{g \text{ water}} \right)$$

(Appendix E in the lab manual has a table listing the density of water at several temperatures.)

Mass. While determining the mass of the sample may seem straightforward (flask/cap/sample – flask/cap), you have to consider the fact that, in Chem 142 lab, we do not have a practical means of evacuating all the air out of the flask. In reality, the calculation for the mass of sample should be written (flask/cap/sample – empty flask/cap).

In the previous discussion about the volume of the flask, we were able to determine the mass of water from (flask/cap/water – flask/cap) because the mass of air in that volume is negligible compared to the mass of the water. In the calculation of the mass of sample, however, the mass of air has the same order of magnitude as the mass of the sample, so (flask/cap/sample – flask/cap) could result in a number close to zero, or even less than zero. To correct this problem, you will determine the mass of air in your flask so that flask/cap is truly just the mass

of the empty flask plus the mass of the cap. To do this, we will find the density of air under experimental conditions for P and T. With the density of air (g/mL) and the volume of the flask, we can calculate the mass of air in the flask.

To calculate the density of air, we will assume the air in the lab is 20.0% oxygen (O₂) and 80.0% nitrogen (N₂). Using the ambient pressure and temperature (there is a thermometer with the barometer on the wall in the lab), a mole-to-volume density can be calculated

$$\frac{n}{V} = \frac{P}{RT}$$

According to Dalton's law of partial pressures (section 5.5 in Zumdahl), we know that total pressure is equal to the sum of the pressures of the individual gases present:

$$P_{total} = P_{O_2} + P_{N_2}$$

If we assume the pressure in the lab is 1.00 atm, then P for O₂ = 0.200 atm and P for N₂ = 0.800 atm. For each gas, we can calculate the mole-to-volume density and then use the molar mass to determine the density in g/mL. Like with Dalton's law of partial pressures, the total density of air is the sum of the densities of the individual gases:

$$\frac{g}{mL} O_2 + \frac{g}{mL} N_2 = \frac{g}{mL} air$$

Having determined the volume of the flask, previously, we can now use that with the total density to calculate the mass of air in the flask.

$$mass\ of\ air\ in\ the\ flask = (mL\ in\ flask) \left(\frac{g\ air}{mL} \right)$$

Now that you have the mass of air, you can subtract that from the mass of the flask and cap so that you have a value for a truly empty flask. You will then be able to determine the mass of your sample:

$$Mass\ of\ sample = (flask/cap/sample - empty\ flask/cap)$$

Safety Considerations

- Take care to properly set up the apparatus so that it is stable. You will need to move quickly to cap the flask once you remove the heat source and an unstable setup could easily tip over, spilling boiling water and shattering glassware all over.

- Once you have capped your flask and remove it from the water bath, the vapor will begin to condense as the flask and gas cool. As the temperature decreases, the pressure will also decrease ($PV = nRT$) resulting in a partial vacuum inside the flask. At this point the flask will be quite fragile and if it breaks, it will implode. Be very careful as you transport it back and forth to the balance.
- The troughs of water the present the possibility of water spills. Be sure to watch for spilled water and clean it up as soon as possible so that no one slips and falls and so that your notebooks and manuals do not get ruined.