Experiment 13: Molecular Weight of Carbon Dioxide

(This experiment is from Santa Monica College, Chemistry 11)

Purpose

The purpose of this experiment is to determine the molecular weight of carbon dioxide, using the ideal gas law. The experimental molecular weight determined will be compared to the known molecular weight of carbon dioxide.

Background

The molecular weight (MW) of a substance is determined by:

$$MW = \frac{\text{grams}}{\text{moles}}$$

The grams can be determined by using the lab balance. The moles can be calculated with the ideal gas law, $PV = nRT$.

Therefore, in this experiment, you will measure the **mass** of the CO_2 sample, as well as the **temperature**, **pressure**, and **volume** of the gas sample.

The carbon dioxide gas is produced by the reaction of calcium carbonate and hydrochloric acid:

$$CaCO_{3\,(s)} + 2\ HCl_{(aq)} \rightarrow CO_{2\,(g)} + CaCl_{2\,(aq)} + H_2O_{(l)}$$

The carbon dioxide gas will be collected in an Erlenmeyer flask that is covered with aluminum foil. The aluminum foil and the fact that carbon dioxide gas is denser than air slows the rate at which the carbon dioxide mixes with the air outside of the flask.

The **mass** of the gas will be determined by mass-by-difference. Be sure to obtain the mass of the empty flask covered by the foil, and then the mass of the foil covered flask filled with carbon dioxide. Before subtracting these two masses to get the mass of the CO_2, the mass of the air in the 'empty' flask must be accounted for.

The **temperature** of the gas will be measured by briefly placing a thermometer underneath the foil, after the final mass of the gas sample is obtained.

The **volume** of the gas will be equal to the volume of the flask. The volume of the flask will be measured by filling it to the brim with tap water, and measuring that water's volume with a graduated cylinder. This is done after the mass and temperature are determined.

The **pressure** of the gas will be equal to the pressure in the room, since the flask is not sealed. The pressure in the room can be read from the barometer.

The mass of the air will be calculated with the volume of the flask and the density of air. See Table 1 for the density of dry air. Assume the air in the lab is dry.

Density = (mass / volume)   mass = (Density · volume)

Table 1. Density of Dry Air

Temperature, °C	**Density, g/L**		
	P = 750 torr	P = 760 torr	P = 770 torr
17	1.201	1.217	1.233
18	1.197	1.213	1.229
19	1.193	1.209	1.225
20	1.189	1.205	1.221
21	1.185	1.201	1.216
22	1.181	1.197	1.212
23	1.177	1.193	1.208
24	1.173	1.189	1.204
25	1.169	1.185	1.200

Information from the CRC Handbook of Chemistry and Physics, 64^{th} ed., 1983-4

Calculate the mass of the dry air, and subtract that value from the mass of the 'empty' flask. This will give the true mass of the 'empty' flask.

$$\text{mass of } CO_2 = (\text{mass}_{\text{CO2 + flask + foils}}) - (\text{mass}_{\text{flask + foils + air}}) - (\text{mass}_{\text{air}})$$

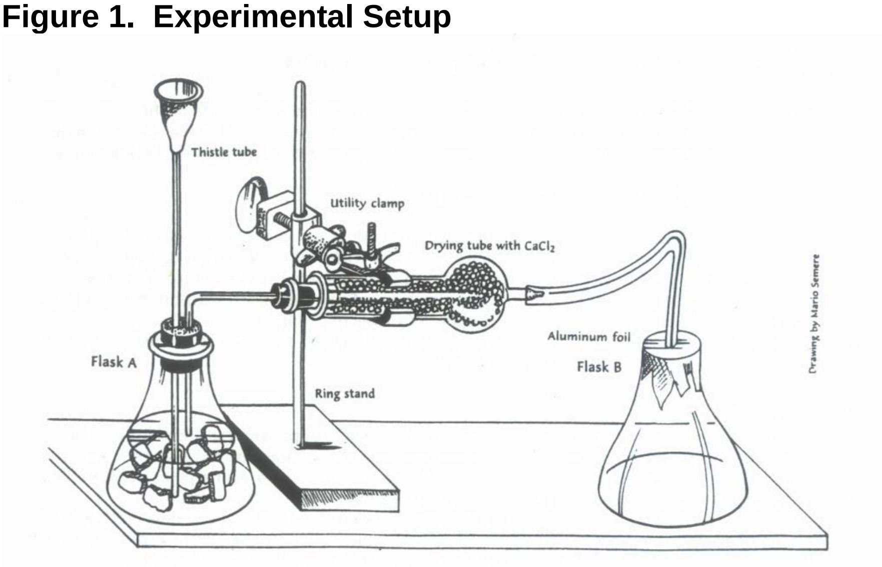
Chemicals

Al foil, approximately 15 cm x 15 cm, 2 pieces (one with a hole and one without a hole)
$CaCO_3$, in the form of marble chips (limestone)
$CaCl_2$, anhydrous solid, for the drying tube
HCl, 6 M *(HCl can cause serious chemical burns and blindness)*
Water, deionized and tap

Equipment

One experimental setup as shown in Figure 1

Figure 1. Experimental Setup



Apparatus for Generation and Collection of Carbon Dioxide Gas

Procedure

1. Obtain one of the 250 mL Erlenmeyer flasks. This will be designated as flask **B**. Make sure that the flask is both clean and dry. Record the mass of the flask and both pieces of foil to the nearest 0.001 g using the balance. Note that it is mass $_{\text{(flask + foils + air)}}$ that is being measured here. Also record the temperature of the air in the flask.

2. Place about 25 grams of marble chips into the other 250 mL Erlenmeyer flask designated as flask **A**. Add enough deionized water to cover the chips. Insert the two-holed rubber stopper (with the thistle tube and the bent tube) into flask *A*, making sure that the thistle tube is adjusted so that it is beneath the water but not touching the bottom of flask. The stopper must fit tightly into the flask. Also obtain about 25 mL of hydrochloric acid (6 M HCl), in a small, labeled beaker.

3. Assemble the apparatus shown in **Figure 1**. Insert the straight glass tube into flask B by placing it between the foil and the flask and pressing the foil against it to hold it in place (use only the Al foil piece with the hole). Be careful to fold and shape the foil only as much as necessary since it is fragile and will easily tear. Attach the flexible end of the tubing to the drying tube. The small rubber stopper attached to the bent glass tubing should be inserted into the other end of the drying tube.

4. When all is ready pour 10 mL of the hydrochloric acid into the top of the thistle tube and allow it to run through the tube and into flask **A**. The reaction should begin immediately as evidenced by gas evolution. Allow the reaction to continue for at least 20 minutes to displace all of the air in flask **B** with carbon dioxide gas. During this time pay attention to what is happening in flask **A**; if the gas evolution ceases, add additional HCl solution through the thistle tube. *(Don't let solution rise up into the stem of the thistle tube; raise the tube as needed.)* After the 20 minutes remove the tube from flask **B** (keep the foil in place) and **immediately cover flask B with the second piece of aluminum foil.** Then weigh the flask containing carbon dioxide on the analytical balance to the nearest 0.001 g.

5. Reassemble the apparatus and allow gas evolution to continue and flow into flask **B** for an additional 15 minutes. Again, weigh flask **B** (with both pieces of foil). The two masses (before and after the 15 minutes) should agree closely (to within 0.005 g). If the mass has increased by more than 0.005 g reassemble the apparatus and continue to collect carbon dioxide gas for an additional 5 minutes then reweigh the flask. Continue this procedure until successive masses agree to within 0.005 g (or until a decrease in mass is observed). Next, measure and record the temperature of the carbon dioxide in the flask. Use the barometer to read the atmospheric pressure.

6. Finally, in order to determine the volume of flask **B**, fill this flask with tap water to the brim. Use the large 500 mL graduated cylinder to measure the volume of water in the flask. Since the flask does not have a pouring lip, place the empty graduated cylinder upside down, over the opening of the flask, and then invert the flask and cylinder in one motion, so the water pours from the flask into the cylinder.

Calculations

1. Calculate the mass of the air in the 'empty' flask $mass_{air}$ = (density · volume)
2. Calculate: mass of CO_2 = ($mass_{CO2 + flask + foils}$) – ($mass_{flask + foils + air}$) – ($mass_{air}$)
3. Calculate the moles of CO_2, using PV = nRT (use correct units)
4. Calculate the molecular weight of CO_2. MW = grams CO_2 / moles CO_2

Show all of your calculations in your notebook.

How well does your experimentally determined MW agree with the known MW?

Calculate the %error, and comment on the %error in your conclusion.