

Lab 3: The molar mass of carbon dioxide

Objectives:

1. Apply the ideal gas law to determine the molar mass of a gaseous compound
2. Think about the design of such an experiment

Skills:

- Set up and execute a multi-step measurement
- Use the Internet as a resource for chemical information

Background:

The **ideal gas law** gives the relationship between the characteristic properties of a gas: $PV = nRT$ where P is the gas **pressure**, V is the gas **volume**, n is the amount (in this case, **moles**) of gas and T is the gas **temperature**.

Of course, the number of moles of a gas is equal to the mass of a given amount of gas divided by the gas's molar mass. To put it another way, the molar mass (MM) of a gas is equal to the mass of a given amount of gas divided by the number of moles in that amount of gas.

Symbolically, $MM = m/n$, where m is the mass of a given amount of gas.

The ideal gas law can be algebraically rearranged so that it is solved for the number of moles of gas: $n = PV/RT$.

Combining the previous two equations:

$$MM = mRT/PV$$

In other words, you can find the molar mass of a pure gaseous substance by measuring the mass of gas in a certain space, the temperature and pressure of the gas and the volume of the space. Recall that R is the ideal gas constant, whose value can be found in nearly any reference (for instance, the inside back cover of the text).

In this experiment, you will fill a flask full of carbon dioxide, weigh it and determine the volume, pressure and temperature of the carbon dioxide and calculate the molar mass of carbon dioxide using the equation above.

One complication is that weighing the carbon dioxide gas is hard; you can't simply tare the balance and add carbon dioxide. You will need to figure out the mass of the air inside the flask and subtract that off of the mass of the flask in order to get just what the flask weighs.

You will need the density of air. You can look up this value in the CRC or James Ierardi's Fire Protection site at Worcester Polytechnic Institute: http://users.wpi.edu/~ierardi/FireTools/air_prop.html. Note that you will need certain input values. Cite the source of the air density you use, and make sure you get

enough significant figures so it is not limiting (you will also have to do a unit conversion). As to the formula that is used by the site to calculate the air density, there is a pretty good Argonne National Labs discussion at <http://www.newton.dep.anl.gov/askasci/wea00/wea00270.htm>

Prelab: Read sections 7.1, 7.2 and 7.8 in the textbook, as well as this handout, to prepare for a **closed-book** prelab quiz before you start the experiment. Also, complete the attached pre-lab calculation.

Materials needed:

- 250 mL Erlenmeyer flask
- Electronic balance
- Barometer (attached to classroom wall)
- Small piece of dry ice
- Rubber stopper
- Thermometer
- 100 mL graduated cylinder

Safety issues: Dry ice sublimates at -78.5°C ; that is the temperature of the dry ice in this experiment. It will cause burns upon prolonged or repeated exposure to bare skin, so **do not pick up the dry ice with your bare hands**; use a paper towel. Carbon dioxide and water are not considered hazardous, so all fluids can be poured down the sink.

Procedure:

1. Obtain the Erlenmeyer flask and rubber stopper; **make sure they are dry**. Push the rubber stopper firmly into the flask (not so firmly that the stopper won't come back out). Weigh this on the balance, and enter the value on the data sheet.
2. Remove the rubber stopper and obtain a small piece of dry ice, about one-third the size of the rubber stopper. Do not handle the dry ice with your bare hands; use a piece of towel paper as shown. Put the dry ice into the flask. Do **not** put the stopper on the flask just yet.
3. Allow the dry ice to sublime **completely**; do not move or shake the flask. After the last bit of dry ice is gone, **wait three minutes**, then firmly stopper the flask and weigh it. Note this value on the data sheet. Note the **temperature** of the room on the data sheet.
4. A 250 mL Erlenmeyer flask is never exactly 250 mL; you will have to determine its volume. Using the flask, water and graduated cylinder, determine the volume of the flask to the nearest tenth of a milliliter. **Take into account the rubber stopper's volume!** Write the volume on the data sheet. You should write down how you did this part (see question 3 in the analysis).
5. Measure the atmospheric pressure in the room by reading the barometer on the wall. If you are not sure how to set the barometer or to read a Vernier scale, please ask the instructor. Note this value on the data sheet.
6. Look up the density of air in g/mL from any resource; write this value on the data sheet and also cite the reference. Calculate the mass of the air in the flask from the information you have, and enter this value on the calculation table.

7. Calculate the mass of the flask alone, then the mass of the carbon dioxide gas in the flask.
8. Finally, calculate the molar mass of carbon dioxide. **Be sure that all of the variables are converted to the correct units.**
9. Repeat this experiment **once more** (clearly you will not need to look up or measure as many things, since most properties will have remained the same).
10. Determine the percent error of your average molar mass, using the reference value of 44.01 g/mol for carbon dioxide. Enter this in the last line of the data sheet.
11. After all of the groups have entered their average molar masses on the transparency, copy the information down and use Excel or your calculator to determine the average for the class, and calculate the percent error for the class off of the accepted value.

Analysis and Questions (Please use your own paper for the answers of these questions, but number them the same way):

1. Show, in excruciating detail, the calculation steps of how you used your data to finally arrive at the calculated molar mass of carbon dioxide.
2. a. Why couldn't you simply tare the flask and weigh the carbon dioxide directly after it sublimed?
b. Recalculate the molar mass of carbon dioxide if you *had* simply tared the flask and let the dry ice sublime.
c. (Tough question) What does this value **represent**?
3. Describe **step-by-step** how you measured the volume of the Erlenmeyer flask.
4. Why was the mass of the original amount of dry ice not important?
5. Why did you wait three minutes after the dry ice sublimed to weigh the flask? How would your calculated molar mass have changed if you'd immediately weighed the flask after the dry ice sublimed? Give a semi-quantitative rationale for your answer.
6. By the way, what are you **assuming** about the behavior of the carbon dioxide?
7. Who was more **accurate** for the molar mass of carbon dioxide, your team or the class? Justify your answer with the percent errors you calculated.

Pre-lab for Lab 3: Molar mass of carbon dioxide

Complete the table for the mystery gas's molar mass (it's not carbon dioxide). You will have to look up the density of air in a table or on the website given in the introduction.

	Trial 1
Mass of flask + air + stopper (g)	223.763
Mass of flask + mystery gas + stopper (g)	223.884
Room air temperature (°C)	27.5
Room air pressure (mm Hg)	768.1
Volume of flask (mL)	260.5
Density of air (g/mL)	
Mass of air in flask (g)	
Mass of flask + stopper (g)	
Mass of mystery gas (g)	
Molar mass of mystery gas (g/mol)	

Data sheet for Lab 3: Molar mass of carbon dioxide

	Trial 1	Trial 2
Mass of flask + air + stopper (g)		
Mass of flask + carbon dioxide + stopper (g)		
Room air temperature (°C)		
Room air pressure (mm Hg)		
Volume of flask (mL)		
Density of air (g/mL)		
Mass of air in flask (g)		
Mass of flask + stopper (g)		
Mass of carbon dioxide (g)		
Molar mass of carbon dioxide (g/mol)		
% error from accepted molar mass of carbon dioxide		
Class average molar mass of carbon dioxide (g/mol)		
% error from accepted molar mass of carbon dioxide		