

Chemistry 150

Please have the purpose and materials ready **before** class on Wednesday, January 24. Write an **abstract** and **paper-clip** it to the front of your individual writeup. The abstract and the carbon-copy pages of the write-up are due in class on **Monday, February 5**.

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.

Your name, your partner's name, date of experiment

Lab 3: Determination of the molar mass of carbon dioxide

Part 1. Purpose

Calculate the molar mass of carbon dioxide gas using the ideal gas law.

List the relevant mathematical equations that you will be using.

Part 2. Materials and methods

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

Equipment needed: **Sketch** the setup and **label** the various pieces of equipment.

Part 3. Procedure

We talked about this in class. Please write down the exact steps you used to arrive at your data (do not list “do the calculation” as a step). Include enough detail (certainly mention all of the glassware you used and when) so that a Chemistry 101 student can get the same results as you did. Also include the criterion for ending the experiment.

Safety issue — Dry ice is extremely cold (-77°C) and will cause frostbite with prolonged contact. Always use gloves or a piece of paper towel to handle dry ice.

Part 4. Original data

If you did not do this already, arrange all of your data in table form, ideally one table for both trials you ran.

Part 5. Calculated results

For one trial, show how you calculated the molar mass of CO_2 , using your data. Pay attention to significant figures.

If you had two trials that resulted in molar masses within 5% of each other, calculate the mean.

Calculate a **percent error** between either your mean (if two good trials) or best value (if only one good trial) for the molar mass of CO_2 and the actual molar mass.

Part 6. Group results

Write your mean value for the carbon dioxide molar mass on the overhead and record **all** groups' values. Calculate the mean and standard deviation, and comment about where your results fit in (for instance, are your results **outliers**? are there any outliers at all in the whole data set?). Finally comment on whether the class as a whole or your partnership was closer to the actual value and why this might be so.

Part 7. Questions

1. 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**?
2. What data value would have changed if you measured the mass of the flask and the sublimed carbon dioxide **immediately** after the dry ice had sublimed? How would your calculated molar mass have changed if you had **not** waited for the gas to reach room temperature?
3. By the way, what are you **assuming** about the behavior of the carbon dioxide?

Part 8. Conclusion

A sentence:

- To report the class's molar mass mean (with standard deviation and proper units) and the percent error between that value and the actual molar mass of carbon dioxide.
- To comment on the size and direction (higher or lower?) of the error and to suggest a source(s) of either systematic or random error, as appropriate.
- To suggest a "fix" for the error source(s) you named.
- Comment on whether you have more confidence in your results or the class's results.

Abstract

Remember to word-process this and put it on a separate page. Attach it to the front of your report

Your name and partner's name, North Seattle Community College
DETERMINATION OF THE MOLAR MASS OF CARBON DIOXIDE USING THE
IDEAL GAS LAW

Using the ideal gas law, we determined the molar mass of carbon dioxide to be _____ (units?), which is _____ % off of the actual value. We have (strong/some/no) confidence in our result because (all obvious sources of error were eliminated/most sources of error were eliminated, except....). For future trials, we would (not change the procedure at all/change the following step(s):.....).