EXPERIMENT 9C

DETERMINATION OF THE IDEAL GAS CONSTANT, R

Problem: Measure the volume, temperature, and number of moles of oxygen gas generated by heating potassium chlorate. Use these values along with the atmospheric pressure measured with a barometer to calculate the ideal gas constant.

FOR PRE-LAB ASSIGNMENT, see page 9C-7

LEARNING OBJECTIVES

1. Proper use of gas-collection apparatus and gas-tight connections
2. Proper use of an Eco-celli barometer
3. Determination of the partial pressure and volume of a gas collected over water
4. Application of the ideal gas law and stoichiometry to experimentally determine R.

MATERIALS

Gas bottle, collection beaker**,** preassembled stoppers and tubing for gas-tight connections as shown in figure 1, ring-stand and clamp, hose clamp, Bunsen burner, glass wool, pipet bulb, potassium chlorate, manganese (IV) oxide.

BACKGROUND

The ideal gas law relates the physical properties of gases through the ideal gas constant (R). The properties accounted for in this law are pressure (P), volume (V), temperature (T) in Kelvin, and amount (moles = n). The amount of any gas is proportional to its pressure and volume, and inversely proportional to temperature.

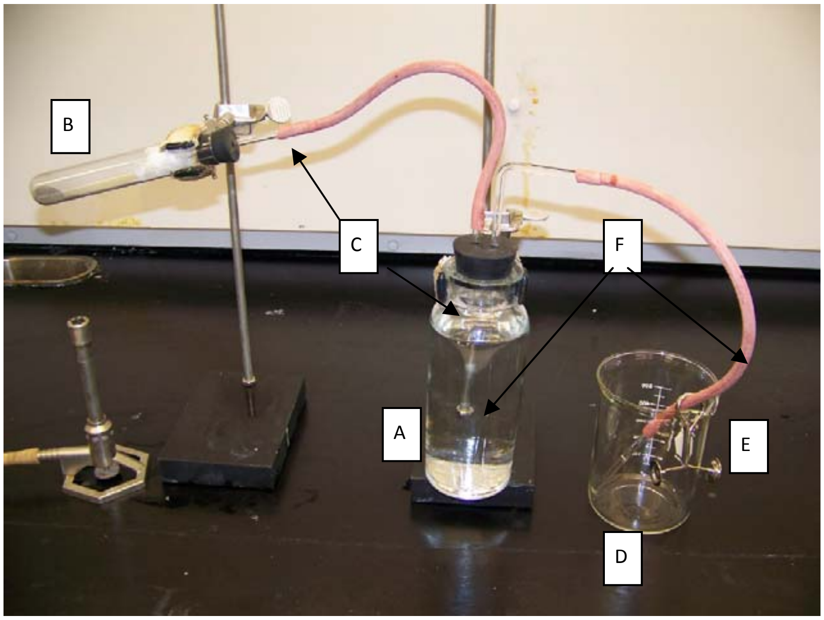
The ideal gas constant R can be found experimentally by determining the number of moles of a gas that occupies a particular measured volume at a known pressure and temperature. This information is determined in this lab by collecting oxygen generated by the decomposition of potassium chlorate over water.

The standard value of R varies depending on the units in which it is given; careful attention to (and understanding of) SI units and unit conversions is essential for success in this lab.

PROCEDURE

*** dispose of ALL waste in the instructor hood unless otherwise instructed***

1. Assemble the apparatus as shown in the following figure (1):



a.) The bottle (A) should be clamped at the neck and the test tube (B) held at a shallow angle and supported by a second ring stand. The bottle and test tub are connected by the gas inlet tubing (C).

b.) The bottle should be filled with water, and the beaker (D) filled with 100 mL of water.

c.) The clamp (E) should be placed on the outlet siphon (F), as close to the end in the collection beaker as is practicable.

d.) Remove the test tube and ensure that it is dry inside, then insert enough potassium chlorate to fill the round bottom of the tube when held vertically.

e.) Mix the potassium chlorate thoroughly with a pinch of manganese (IV) oxide (MnO_2), which acts as a catalyst.

f.) Put a plug of glass wool into the mouth of the tube and weigh the tube and its contents on an ***analytical*** balance.

g.) Before replacing the test tube on the apparatus, open the pinch clamp (E) on the outlet siphon (F) and prime the siphon by forcing air into the inlet tube (C) *with a pipet bulb* until the entire outlet siphon tube is filled with water, then close the clamp to keep it primed.

h.) Replace the test tube on the rubber stopper with a ***gentle twisting motion*** to insure a safe and gas-tight connection.

i.) With the outlet tube opening *under* the beaker water, raise the beaker so that the water in the beaker and in the bottle are at the same level, then open the pinch clamp. If the pressure in the bottle is greater than that of the atmosphere, water will flow from the bottle into the beaker. If the internal pressure is lower, water will be forced into the bottle by the atmosphere. After time to equilibrate (about 15 seconds) close the pinch clamp. The air in the test tube and the bottle is now at the prevailing barometric pressure.

j.) Empty the collecting beaker and replace it before proceeding.

k.) **Get instructor approval of your apparatus before proceeding.**

Instructor's initials___________________

2. Open the pinch clamp. A few drops of water should run into the beaker, but no more. If a significant amount of water escapes, you have a leak. Reset the apparatus and try again. Possible problems: the pinch clamp may be too far from the end of the siphon, the apparatus may not be effectively gas-tight, or the pressure in the bottle was not adjusted properly.

3. Tap the tube containing potassium chlorate in order to distribute it along the side of the tube.

4. With the pinch clamp open, heat ***gently*** with a ***small*** flame that is moved back and forth under the test tube, gradually proceeding from the bottom toward the top of the tube as the chlorate melts and decomposes (Do not allow the rubber stopper to come in contact with the molten potassium chlorate or the flame!). Keep the rate of heating just sufficient to produce a steady flow of water from the nozzle into the beaker. If heating is too rapid, pressure will build up in the bottle and blow out the stopper. You do not need completely fill the collection beaker with water, but collect at least 300. mL of water. Stop the heating when the water level in the bottle is ***no lower*** than two inches above the siphon inlet in the bottle.

5. With the pinch clamp ***open*** and the nozzle below the water level in the beaker, let the apparatus cool to room temperature. Keep the pinch clamp open and make sure the nozzle remains below the water level in the beaker while the test tube cools. (The test tube should feel as cool to the touch as the bottle before you can proceed to step 7.)

6. While the apparatus is cooling, read the barometer and measure the temperature of the water in the beaker.

7. When the test tube has cooled, note the difference in the relative water levels in the bottle and the beaker. Slowly raise or lower the beaker (keeping the nozzle under water) until the level of water in it is the same as that in the bottle; then close the pinch clamp. The mixture of oxygen and air in the test tube and the bottle is now at the prevailing barometric pressure.

8. Measure the volume of the water collected (equal to the volume of gas produced) using your graduated cylinder.

9. Weigh the test tube and its contents once again to find the weight of the oxygen evolved.

10. Repeat steps 1-9 again for a second set of results.

EXPERIMENTAL DATA

Observation:	Data Set 1:	Data Set 2:
Mass of tube with potassium chlorate and manganese (IV) oxide	___________	___________
Mass of tube and reaction residue after water has been collected	___________	___________
Volume of water collected	___________	___________
Temperature of water collected	___________	___________
Atmospheric pressure	___________	___________
Partial pressure of H_2O [1]	___________	___________

[1] Values for the vapor pressure of water can be found in the CRC Handbook, on a placard near the barometer in your laboratory, or on the Plebe Resources website (under Chemical Safety Information).

DATA TREATMENT

1. Calculate the partial pressure of oxygen in the bottle for *each set of data*.

2. Calculate the moles of oxygen in the bottle for *each set of data*.

3. Calculate the value of R from *each set of data*, in L•atm•mol^{-1}•K^{-1}. Show your work!

4. Average the R value of the two data sets.

QUESTIONS

1. Calculate the % error of your average value of R from the literature value in the textbook.

2. The ideal gas constant, R, can be calculated and expressed with a variety of units. Convert your average value of R to the thermodynamic form, $J \bullet mol^{-1} \bullet K^{-1}$ (realize that this is equivalent to $L \bullet kPa \bullet mol^{-1} \bullet K^{-1}$, and 1 atm = 101.325 kPa).

3. How would your determination of R be affected if all the potassium chlorate does not decompose? Will your value be too high, too low, or unchanged? Explain your answer.

4. If a student neglects to equilibrate the fluid levels in the bottle and collecting beaker ***after*** the reaction has stopped, will this affect the determination of R? If so, in what way? Explain your answer.

PRE-LAB ASSIGNMENT

1. What is the mathematical expression of the ideal gas law?

2. Balance the following equation: ____ $KClO_3\ (s) \xrightarrow{\Delta}$ ____ $KCl\ (s)$ + ____ $O_2\ (g)$

3. Carefully read the experimental procedure and examine figure 1. Then answer these questions based on the following data. A midshipman running this experiment found that the initial weight of the test tube and potassium chlorate was 30.800 grams. After heating, the mass of the tube was 30.500 grams.

a) What materials remain in the test tube?

b) What product(s) are found in the bottle?

c) What is the mass of the oxygen in the bottle?

d) How many moles of oxygen were produced?

e) What other gases are in the bottle? (remember that oxygen was bubbled through water)

f) How is the partial pressure of oxygen in the bottle related to the atmospheric pressure in the room, once the bottle and beaker water levels are matched (step 7)?

g) What is the relationship between volume of oxygen generated and volume of water collected in the beaker (pushed out of the bottle)?