MOLAR MASS OF OXYGEN DETERMINED EXPERIMENTALLY

Reminder – Goggles must be worn at all times in the lab!

PRE-LAB DISCUSSION:

In this particular lab, we are going to try to determine the MOLAR MASS of OXYGEN experimentally. It is common knowledge that the molar mass of O_2 is 32.00 grams per mole, but you are going to see how close you can come to this in the lab. From our studies of the gas laws, we know that $PV = nRT$ which can be rearranged to solve for n, the number of moles:

$$n = PV/RT$$

We are trying to find the molar mass, which has units of grams/mole. We will measure the mass of the oxygen generated by the potassium chlorate indirectly, and we will measure the volume of the gas at room conditions. We merely need to plug our conditions of volume, temperature, and pressure into the ideal gas law in order to find n, the number of moles of oxygen collected. Once we know the mass, and the number of moles, we will divide the mass, g, by the number of moles, n.

$KClO_3$ decomposes slowly when heated. In order to speed up this reaction, we will be adding manganese dioxide (MnO_2) as a CATALYST.

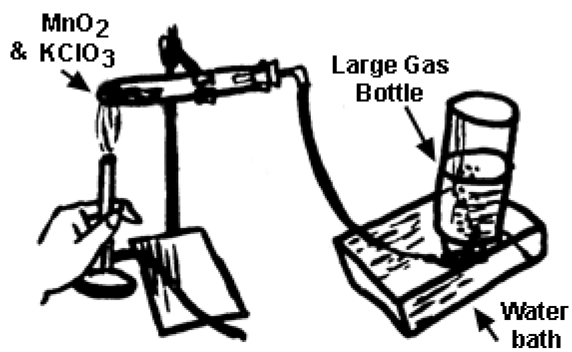
$$2KClO_3 \xrightarrow{\Delta, MnO_2} 2KCl + 3O_2$$

PURPOSE:

To experimentally determine the molar mass of oxygen.

PROCEDURE:

1. Weigh out approximately 0.75 grams of potassium chlorate into your large ignition tube. Now add approximately 0.75 grams of manganese dioxide to the same tube.
2. With the stirring rod, mix the two chemicals together well. Next, weigh the ignition tube and its contents accurately, and record the mass in your Data section. Remember to record the mass to an appropriate number of places to the right of the decimal.
3. Set up the apparatus for collecting a gas by water displacement as shown to the right. Clamp the ignition tube in place and connect the delivery tube. Make sure the stopper fits snugly in the opening of the ignition tube so that no oxygen can escape around the sides. Completely fill the bottle with water.



4. Begin heating the tube and collecting the oxygen. Continue heating until no more oxygen is liberated, or until the water level in the collecting bottle gets near the bottom. Turn off the Bunsen burner when the reaction is completed.
5. Remove the delivery tube from the water immediately after the oxygen has stopped generating. If you do not, water will be pushed back into the test tube as the tube cools (due to greater outside pressure), and the experiment will have to be repeated. At the same time, remove the tube where it vents into the gas collecting bottle. If you do not, the atmospheric pressure will push the remainder of the water out of the gas collecting bottle, and you will have to start over.
6. When the ignition tube has cooled, weigh the tube and its contents and record the mass in the Data section.
7. Using a 100 ml graduated cylinder, measure the volume of the gas collected in the bottle. This may be done indirectly by:
   a) Measuring the water left in the gas-collecting bottle with a 100 mL graduate.
   b) Measure the total volume of the bottle by completely filling the bottle to the top with water, and measuring that volume out into a 100 mL graduated cylinder.
   c) Subtracting the volume of water remaining after the lab from the total volume of the gas-collecting bottle. Record the oxygen volume in the Data section.
8. Using a thermometer, take the temperature of the water in the trough and record the value.
9. Read the barometric pressure in the lab in mm Hg (torr) and record this value.

RESULTS:

Observations and Data

Mass

1. Mass of ignition tube and contents (before heating) ________________grams
2. Mass of ignition tube and contents (after heating) ________________grams
3. **M** = Mass of oxygen collected (line 1 – line 2) ________________grams

Volume

4. TOTAL volume of the gas collecting bottle ________________mL
5. Volume of water remaining after gas collection ________________mL
6. Volume of oxygen collected (line 4 – line 5) _______________mL
7. **V** = Volume of oxygen, converted to liters (line 6 ÷ 1000) ________________L

Temperature

8. Temperature of water bath in °C _______________°C
9. **T** = Temperature of water in K ________________K

Pressure

10. Barometric pressure in mm Hg ________________mm Hg
11. Vapor pressure of water at the above temperature ________________mm Hg
    (see vapor pressure chart on page 3)
12. Pressure of dry oxygen (line 10 – line 11) ________________mm Hg
13. **P** = Pressure of dry oxygen in atmospheres (line 12 ÷ 760) ________________atm

Calculations (Show your work! Observe significant figures and include units.)

1. Calculate the number of moles of oxygen collected, using $n = \frac{PV}{RT}$
2. Calculate the experimental molar mass of O_2 by dividing the mass(**M**) by *n*, the number of moles.
3. Calculate the ABSOLUTE ERROR, the difference between the accepted value and your experimental value.
4. Calculate the percentage error in your result.

Water Vapor Pressure Table

Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)
0.0	4.6	19.5	17.0	27.0	26.7
5.0	6.5	20.0	17.5	28.0	28.3
10.0	9.2	20.5	18.1	29.0	30.0
12.5	10.9	21.0	18.6	30.0	31.8
15.0	12.8	21.5	19.2	35.0	42.2
15.5	13.2	22.0	19.8	40.0	55.3
16.0	13.6	22.5	20.4	50.0	92.5
16.5	14.1	23.0	21.1	60.0	149.4
17.0	14.5	23.5	21.7	70.0	233.7
17.5	15.0	24.0	22.4	80.0	355.1
18.0	15.5	24.5	23.1	90.0	525.8
18.5	16.0	25.0	23.8	95.0	633.9
19.0	16.5	26.0	25.2	100.0	760.0