Name ___________________ **Chemistry** **___/___/___**

Molar Volume of a Gas Lab

Equipment & Chemicals:

Equipment		**Chemicals Used**
50 mL gas-measuring tube	ring stand	magnesium ribbon
one-hole stopper	utility clamp	3 M HCl
400 mL beaker	cotton thread	
10 mL graduated cylinder	2.0 L graduated cylinder (shared)	

Procedure:

☐ 1. Obtain a piece of magnesium ribbon. Record the mass using the analytical balance.

☐ 2. Obtain a piece of cotton thread about 8 cm long. Tie one end of the thread around the piece of magnesium ribbon, leaving a few cm of thread free. Bend the piece of magnesium so that it will easily fit into the gas-measuring tube.

☐ 3. Add 300 mL of water at room temperature to a 400 mL beaker.

☐ 4. Obtain approximately 10. mL of 3 M HCl. Add the HCl to the gas measuring tube.

☐ 5. Using a beaker, **completely** fill the gas measuring tube with room temperature water.

☐ 6. Lower the piece of magnesium ribbon 4 or 5 cm into the gas-measuring tube. Drape the thread over the edge of the tube and insert the one-hole rubber stopper into the tube.

☐ 7. Place your finger over the hole in the rubber stopper and invert the gas-measuring tube. Lower the stoppered end of the tube into the beaker of water. Clamp the tube in place so that the stoppered end is a few centimeters above the bottom of the beaker. Record your observations.

☐ 8. While waiting for your reaction to complete, record the room temperature and the barometric pressure in the data table.

☐ 9. Let the apparatus stand exactly 5 minutes **after the magnesium has completely reacted**. Tap the sides of the gas collecting tube to dislodge any gas bubbles that may have become attached to the sides of the tube. Place your finger over the hole in the stopper and transfer the tube to a 2.0 L graduated cylinder filled with water and remove your finger from the hole.

☐ 10. Move the tube up and down to equalize pressure (the pressure inside the tube with the atmospheric pressure) until the water level in the tube is the same as that in the 2.0 L graduated cylinder. On the scale of the gas-measuring tube, read the volume of the gases in the tube. Record the volume in your data table.

Post-Lab Questions: You must show all work to receive credit.
Unless stated otherwise, use data calculated in early steps to help you answer later problems.

1. Magnesium metal will be placed in an HCl solution. Write the balance chemical equation for this reaction.
2. If 0.3355 g Mg of magnesium are reacted, how many moles of H_2 would be produced?
3. Assuming the reaction is carried out at STP, which it won't be – zero Celsius is too cold to do experiments in a classroom, what volume of hydrogen would be theoretically produced? Use the amount of moles calculated in #2.
4. More realistically, assume the reaction is carried out at 23.0°C and 29.72 in Hg (as read from the classroom barometer), what is the pressure (in atm) of the dry gas?
5. Using the pressure of the dry gas calculated in #4 and the moles calculated in #2, what volume of gas would be theoretically produced at 23.0°C and 29.72 in Hg?
6. **At the above temperature and pressures**, the amount of gas produced will be too much to collect in our 50. mL gas collecting tubes. Determine the maximum mass of magnesium that can be used so that no more than 40. mL of hydrogen gas is produced.
7. Using your results and info used in #5. Calculate the volume of gas produced by 1.0 mole of H_2 gas at STP.

Data Table:

	Trial 1
Mass of magnesium ribbon, g	
Moles of Mg reacted, mol	
Moles of H_2 produced (stoichiometrically), mol	
Temperature of room, °C	
Temperature of room, K	
Barometric Pressure, in Hg	
Barometric Pressure, mm Hg	
Barometric Pressure, kPa	
Vapor Pressure of Water, kPa	
Pressure of the dry gas, kPa	
Pressure of the dry gas, atm	
Volume of H_2 gas in the tube, mL	
Volume of H_2 gas in the tube, L	
Calculated Experimental Volume of 1 mol of H_2 at STP, L	
Theoretical Value = 22.4 Percent Error, %	

Observations & Conclusions:

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Vapor Pressure of Water

Temperature (°C)	Pressure (kPa)
10	1.2281
11	1.3129
12	1.4027
13	1.4979
14	1.5988
15	1.7056
16	1.8185
17	1.938
18	2.0644
19	2.1978
20	2.3388
21	2.4877
22	2.6447
23	2.8104
24	2.985
25	3.169
26	3.3629
27	3.567
28	3.7818
29	4.0078
30	4.2455
31	4.4953
32	4.7578
33	5.0335
34	5.3229
35	5.6267
36	5.9453
37	6.2795
38	6.6398
39	6.9969
40	7.3814
41	7.784
42	8.2054
43	8.6463
44	9.1075
45	9.5895
46	10.094
47	10.62
48	11.171
49	11.745
50	12.344

Post-Lab Questions:

Unless stated otherwise, use data calculated in early steps to help you answer later problems.

1. Magnesium metal will be placed in an HCl solution. Write the balance chemical equation for this reaction.

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

2. If 0.3355 g Mg of magnesium are reacted, how many moles of H_2 would be produced?

$0.3355 \div 24.3 = 0.0138$ mol Mg = 0.0138 mol H_2

3. Assuming the reaction is carried out at STP, which it won't be – zero Celsius is too cold to do experiments in a classroom, what volume of hydrogen would be theoretically produced? Use the amount of moles calculated in #2.

0.309 L

4. More realistically, assume the reaction is carried out at 23.0°C and 29.72 in Hg (as read from the classroom barometer), what is the pressure (in atm) of the dry gas?

29.72 in Hg → 754.888 mm Hg → 100.643 kPa

100.643 kPa – 2.8104 kPa = 97.83 kPa

97.83 kPa → 0.966 atm

5. Using the pressure of the dry gas calculated in #4 and the moles calculated in #2, what volume of gas would be theoretically produced at 23.0°C and 29.72 in Hg?

PV= nRT

V = (0.0138)(0.08206)(296) / (0.966)

V = 0.347 L

6. **At the above temperature and pressures**, the amount of gas produced will be too much to collect in our 50. mL gas collecting tubes. Determine the maximum mass of magnesium that can be used so that no more than 40. mL of hydrogen gas is produced.

n = (.966)(0.040) / (0.08206)(296)

n = 0.00159 moles

0.039 grams

7. Using your results and info used in #5. Calculate the volume of gas produced by 1.0 mole of H_2 gas at STP.

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

$$\frac{(0.966)(0.347)}{(0.0138)(296)} = \frac{(1)(V)}{(1)(273)}$$

V = 22.4

Teacher set-up notes:

The maximum mass of magnesium to be used in 0.0390g. Students must use analytical balance.

Water used to fill the tube, beaker and battery jar must be at room temperature. Fill (12) 1.0 L beakers and the battery jar the day before the lab and let them sit at the lab stations.

The goal of the law is to determine the volume of one mole of a gas. The gas is hydrogen.

$$\frac{P_1V_1}{n_1R_1T_1} = \frac{P_2V_2}{n_2R_2T_2}$$

R is the same for both so it can be eliminated

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

P_1: pressure of dry gas
V_1: volume of gas produced
n_1: moles of gas used
T_1: temperature of room

P_2: standard pressure, 760 mm Hg
V_2: calculated volume of one mole of a gas
n_2: one mole
T_2: standard temperature, 273 K