LAB 8 – The Gas Constant, R

OBJECTIVE: Experimentally determine the gas constant.

BACKGROUND

The ideal gas law is a mathematical description of how the pressure, volume, amount and temperature of an ideal gas relate to each other. The ideal gas law is typically stated as

$$PV = nRT \qquad \text{(Eq. 1)}$$

P = Pressure (atm)

V = Volume (L)

n = Amount of gas (moles)

R = “the gas constant” = 0.08206 L·atm/K·mol

T = Temperature (K)

In this experiment, we will produce hydrogen gas by reacting magnesium with excess hydrochloric acid.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g) \qquad \text{(Eq. 2)}$$

We will determine P, V, n and T and then use those values to calculate R, using a rearranged statement of the ideal gas law.

$$R = \frac{PV}{nT} \qquad \text{(Eq. 3)}$$

Pressure will be determined by a local barometric reading. Volume will be determined by water displacement using the apparatus in Figure 1. Amount will be determined by stoichiometry. Temperature will be approximated by measuring the temperature of the water in the filter flask(b) of the apparatus in Figure 1.

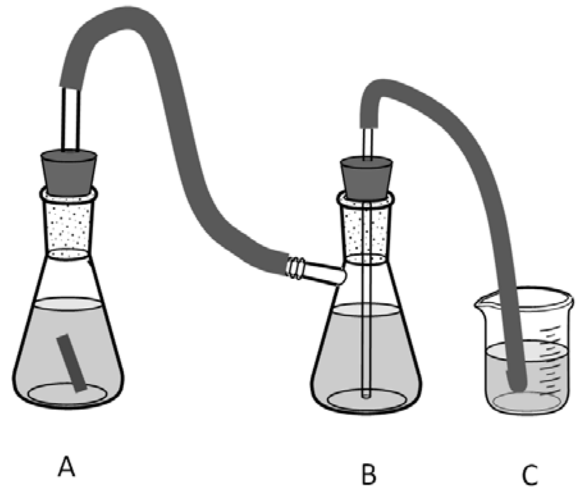

Figure 1 – Experimental apparatus for measuring the volume of H_2 gas produced by the reaction of Mg and HCl.

a) The **“reaction flask”** contains a strip of magnesium and excess HCl. The reaction produces H_2 gas. The reaction is complete when the magnesium is completely consumed.

b) The **“filter flask”** is filled with tap water to just below the side arm of the flask. This water will get displaced by the H_2 gas that is produced in the reaction flask. The temperature of the water at the end of the reaction is used as an approximation of the temperature of the gas.

c) The **“collection beaker”** (initially empty) catches the water that is displaced from the filter flask. The initial and final mass of the beaker is measured, recorded and used to determine the volume of hydrogen gas produced by the reaction.

PROCEDURE

1. Set up the apparatus shown in Figure 1.
2. Fill the filter flask with tap water to a level about a half inch below the side arm of the flask. Do not over fill.
3. Add *about* 30 mL of 1 M HCl to the reaction flask.
4. Obtain a strip of magnesium with a mass of 0.10 to 0.15 g. Measure the mass of the strip (m).
5. Measure the initial mass (m_i) of the empty, dry collection beaker. **Use the high capacity (1500g) balance**.
6. Remove the stopper from the reaction flask and drop in the strip of magnesium. Quickly and carefully replace the stopper. The magnesium will float on the surface of the acid and bubble vigorously. The gas that is produced in the reaction flask will displace water from the filter flask into the collection flask.
7. Wait for the magnesium to be completely consumed by the acid.
8. Measure the final mass (m_f)of the collection flask with the displaced water. **Use the high capacity (1500g) balance**.
9. Measure the temperature (T) of the water remaining in the filter flask.
10. Discard the contents of the **reaction flask** in the aqueous waste container.
11. Repeat Steps 2-10 for a total of 3 runs.

CALCULATIONS

1. Go to a reliable weather app, look up and record the current barometric pressure. It will be in inches of Hg, so you will need to convert it to atm.

$$P_{(atm)} = P_{(in.Hg)} \times \frac{0.0334\ atm}{1\ in.Hg}$$

2. Calculate the volume of gas produced by the reaction.

$$V = (m_f - m_i) \times \frac{1mL}{1g} \times \frac{1L}{1000mL}$$

3. Calculate the amount of gas produced.

$$n = moles\ of\ H_2 = \frac{(\qquad\qquad)gMg}{1} \times \frac{1molMg}{24.3gMg} \times \frac{1mol\ H_2}{1mol\ Mg}$$

4. Convert the approximate temperature of the gas produced to units of Kelvin.

$$T_{Kelvin} = T_{Celcius} + 273.15$$

5. Calculate the value of R, based on the values of P, V, n and T from above, using R = PV/nT.

Table 1 – Experimental data and results

#	m_i Initial mass of collection beaker (g)	m_f Final mass of collection beaker (g)	m Mass of magnesium strip (g)	P Pressure of gas (atm)	V Volume of gas (L)	n Amount of gas (moles)	T Approximate temperature of gas (K)	R Calculated Gas Constant (L·atm/ K·mol)
1								
2								
3								
							Average =	
							S =	

Does your result (avg ± s) agree with the accepted value (0.08206 L·atm/ K·mol)? _______________

What kind of experimental errors could have contributed to the difference between your value and the published value?