Purpose

To determine the value of the ideal gas law, R.

Learning Objectives

Conduct basic measurements, mathematical calculations and conversions for measurement of volume, pressure, and temperature; calculate percent error and account for sources of error in an experimental result.

Conduct laboratory experiments for a redox reaction by using equipment identifiable by name and by following safety procedures.

Perform chemical calculations using mass, mole and volumes and identify limitations of measuring devices in order to state the uncertainty/significant figures in measurements and calculation results.

Demonstrate a proficiency in relating the chemistry of gases to everyday phenomena.

Introduction

Molecules in gases have very little attraction for one another, because relative to their size, they are so far apart from each other. For this reason, we can often assume that a gas behaves "ideally," meaning that the physical behavior of the gas is independent of the composition of the gas. If a gas behaves ideally, then its physical properties can be described by the Ideal Gas Equation (Equation IG.1).

$$P \cdot V = n \cdot \mathrm{R} \cdot T \qquad \text{(Equation IG.1)}$$

In Equation IG.1, P is the pressure of the gas, V is its volume, n represents the number of moles of gas, T is the temperature and **R** is the gas constant. Let's consider why the variables in the ideal gas equation have the relationships expressed in Equation IG.1. The ideal gas equation assumes that there are no attractive forces between gas molecules, similar to two billiard balls bouncing off one another in a totally elastic collision. When gas molecules are placed in a balloon, the collisions occur not only with each other but also with the wall of the balloon, thus causing the balloon to expand. This expansion can be measured as a volume (V). More molecules will cause more collisions and thus more expansion. Therefore, the quantity of molecules or number of moles (n) is directly proportional to the volume.

Another factor to consider is the kinetic energy of these molecules. Kinetic energy is related to the temperature

(T) of the molecules, higher temperatures are associated with greater kinetic energies. If the collisions with the wall of the balloon are happening at a very high speed due to high kinetic energy, the collisions will be harder and more frequent and thus the wall of the balloon will be pushed out further (a larger volume) than if the kinetic energy were less. The result is that volume is also directly proportional to the temperature. This relationship is seen in Equation IG.1 with both n and T on the right side of the equation and V on the left.

Consider what happens when you open the balloon, as the air streams out, the pressure is dropping but the volume that the air is now filling (i.e., the surroundings) goes up implying that volume is inversely proportional to the pressure. (*Note that (P) and (V) are both on the same side of Equation IG.1.)* .) If a captured gas's volume is adjusted such that its pressure is equal to the atmospheric pressure (P) which can be measured with a barometer, this pressure can be related to the other variables, (n), (V) and (T) by a proportionality constant. This constant is called the ideal gas constant (R) and in this experiment we will determine its value. Equation IG.1 can be rearranged so that the ideal gas constant is on one side of the equation by itself (Equation IG.2).

$$R = \frac{P \cdot V}{n \cdot T} \qquad \text{(Equation IG.2)}$$

The accepted literature value of the gas constant, R, is 0.0821 $\frac{\mathrm{L{\cdot}atm}}{\mathrm{mol{\cdot}K}}$. *Note the units for* R because you must use the correct units of the variables in the calculation for R.

Reactive metals undergo reactions with aqueous acids to produce hydrogen gas and metal salts. In these reactions (Reaction IG.1), the metals are oxidized (lose their outer shell electrons)

$$\mathrm{M^0(s) \longrightarrow M^{\mathit{n}+}(aq) + \mathit{n}\,e^-} \qquad \text{(Reaction IG.1)}$$

to form monatomic cations while the hydrogen cations (H^+) from the acid are reduced (gain electrons) to form elemental diatomic hydrogen gas (H-H) as shown in Reaction IG.2.

$$\mathrm{2\,H^+(aq) + 2\,e^- \longrightarrow H_2(g)} \qquad \text{(Reaction IG.2)}$$

If the acid is in excess, **the number of moles of gas produced can be predicted from the mass of the metal, according to the stoichiometric equation for the reaction.** The gas can be collected, and its volume determined under measured conditions of pressure and temperature. From this information it is possible to determine a value for the gas constant, R, as you will see in the sample calculation below.

The metal you will use is magnesium (Mg). The balanced equation for the reaction with HCl is given in Reaction IG.3:

$$\mathrm{Mg^0(s) + 2\,HCl(aq) \longrightarrow H_2(g) + MgCl_2(aq)} \qquad \text{(Reaction IG.3)}$$

The apparatus you will be using to collect the hydrogen evolved is called a *eudiometer*, which looks like a very long graduated test tube. When it stops expanding, the gas bubble formed in the eudiometer will have the same total pressure of gases as the external atmospheric pressure, P_{bar} (measured with a barometer), once you place the eudiometer in the Equalization Chamber, which looks like an overgrown plastic graduated cylinder. The gas will be collected over water. For this reason, the total gas pressure in the bubble collected will be not only from the H_2 produced but will be partially due to the pressure of water vapor ($P_{\mathrm{H_2O}}$), according to Dalton's Law of Partial Pressures (Equation IG.3).

$$P_{\mathrm{total}} = P_{\mathrm{bar}} = P_{\mathrm{H_2}} + P_{\mathrm{water}} \qquad \text{(Equation IG.3)}$$

To determine the pressure of the hydrogen gas, $P_{\mathrm{H_2}}$, the vapor pressure must be subtracted from the atmospheric pressure (read from the barometer), as shown in Equation IG.4.

$$P_{\mathrm{H_2}} = P_{\mathrm{bar}} - P_{\mathrm{H_2O}} \qquad \text{(Equation IG.4)}$$

The vapor pressure of water varies with temperature and can be obtained from the plot of pressure versus temperature (Figure IG.1).

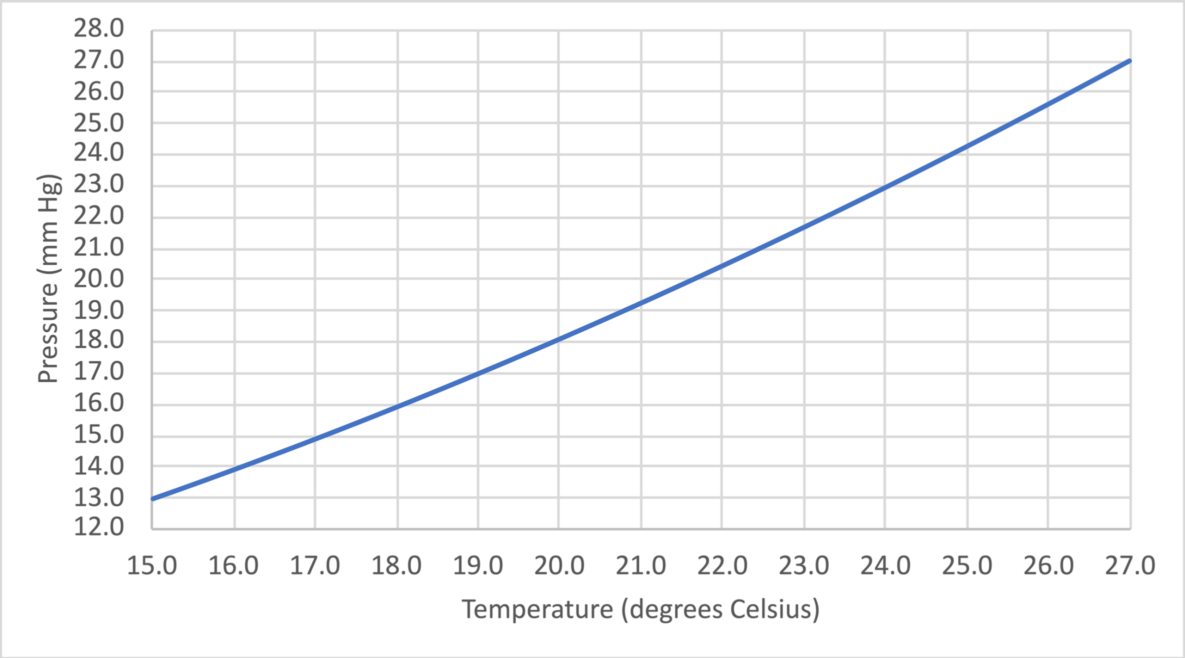

Figure IG.1: Vapor pressure of water over water versus temperature.

Example IG.1

Suppose 0.064 g of Al reacts with excess HCl according to Reaction IG.4.

$$2\,\mathrm{Al(s)} + 6\,\mathrm{HCl(aq)} \longrightarrow 3\,\mathrm{H_2(g)} + 2\,\mathrm{AlCl_3(aq)} \qquad \text{(Reaction IG.4)}$$

If 80.22 mL of gas are collected at 25.0 °C under a barometric pressure/atmospheric pressure of 752.4 mmHg (millimeters of mercury), what value of R will be calculated?
Start by calculating the number of moles of hydrogen formed:

$$0.064\ \mathrm{gAl} \times \frac{1\ \mathrm{mol\ Al}}{26.98\ \mathrm{g\ Al}} = 2.4 \times 10^{-3}\ \mathrm{mol\ Al}$$
$$2.4 \times 10^{-3}\ \mathrm{mol\ Al} \times \frac{3\ \mathrm{mol\ H_2}}{2\ \mathrm{mol\ Al}} = 3.6 \times 10^{-3}\ \mathrm{mol\ H_2}$$

Next, use the vapor pressure graph (Figure IG.1) and the rearranged form of Dalton's Law of Partial Pressures (Equation IG.4) to determine the pressure contributed by the hydrogen gas:

$$P_{\mathrm{H_2O}} \text{ at } 25\ ^\circ\mathrm{C} = 23.2\ \mathrm{mmHg} \text{ (from Figure IG.1)}$$
$$P_{\mathrm{H_2}} = P_{\mathrm{bar}} - P_{\mathrm{H_2O}} = 752.4\ \mathrm{mmHg} - 23.2\ \mathrm{mmHg}$$
$$P_{\mathrm{H_2}} = 729.2\ \mathrm{mmHg}$$

Convert temperature to K, volume to L, and pressure to atm to be consistent with compared literature value.

$$25.0^\circ C + 273.15 = 298.2\ \mathrm{K} \text{ (to the correct number of significant figures)}$$
$$729.2\ \mathrm{mmHg} \times \frac{1\ \mathrm{atm}}{760\ \mathrm{mmHg}} = 0.9595\ \mathrm{atm} \text{ (760 is an exact number)}$$
$$80.22\ \mathrm{mL} \times \frac{1\ \mathrm{L}}{1000\ \mathrm{mL}} = 0.08022\ \mathrm{L} \text{(1000 is an exact number)}$$

Finally, use Equation IG.2 to calculate the value of R:

$$R = \frac{P \cdot V}{n \cdot T} = \frac{0.9595\ \mathrm{atm} \times 0.08022\ \mathrm{L}}{(3.6 \times 10^{-3}\ \mathrm{moles})(298.2\ \mathrm{K})} = 0.072\ \frac{\mathrm{L} \cdot \mathrm{atm}}{\mathrm{mol} \cdot \mathrm{K}}$$

The experimental value for R obtained from these data, then, is $R = 0.072\ \frac{\mathrm{L \cdot atm}}{\mathrm{mol \cdot K}}$.

Note that the experimental value in Example IG.1 differs from the accepted literature value of 0.08206 $\frac{\mathrm{L \cdot atm}}{\mathrm{mol \cdot K}}$. The percent error in this determination (using the literature value as the accepted value) is calculated as shown below.

The straight lines mean absolute value, so if you get a negative value inside those straight lines, report the positive value.

$$\%\ \text{error} = \frac{|\,accepted - experimental\,|}{accepted} \times 100$$

$$\%\ \text{error} = \frac{\left|0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} - 0.072\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right|}{0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}} \times 100 = 13$$

You will use the conversion definition, 760 mmHg = 1 atmosphere to convert from mmHg to atmosphere. This conversion does not define significant figures.

Procedure

1. Prepare a buret stand with a buret clamp.

2. Obtain a sample of magnesium metal and record the mass. **Do not exceed a mass of 0.046 g**. Fold and twist the piece of magnesium metal in a copper wire cage which is attached through the narrow end of a rubber stopper (Figure IG.2a). The stopper with cage is already prepared for you. The cage does not have to totally cover the magnesium but merely support it without the piece of metal falling out of the cage.

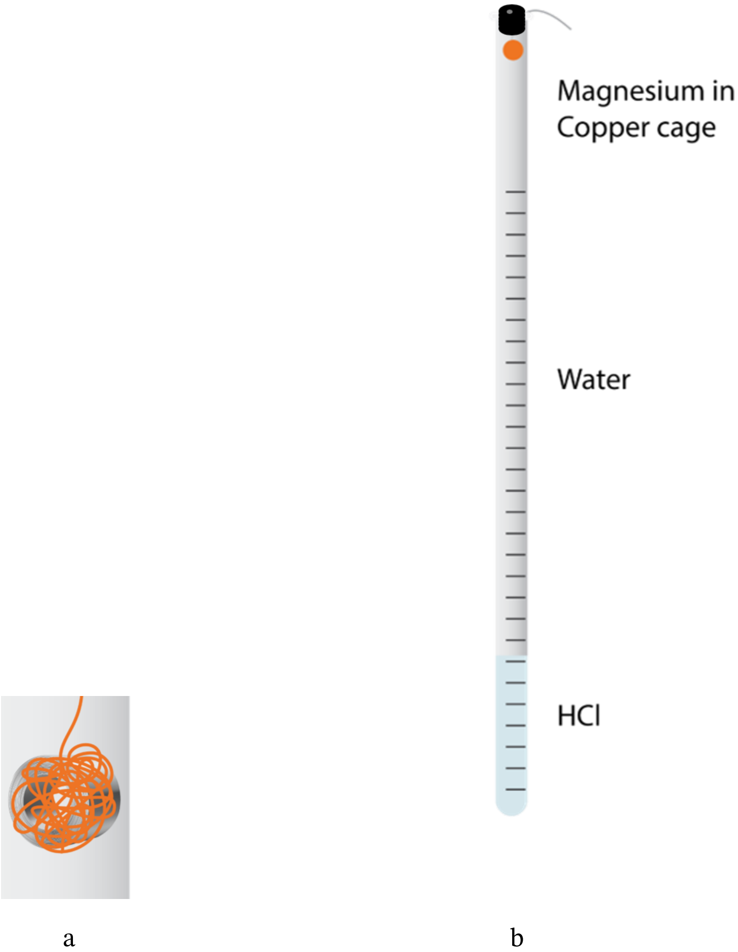

Figure IG.2: a) Mg in the copper cage b) Eudiometer tube showing HCl, water and the Mg in a copper cage.

3. Place the rubber stopper in the eudiometer to check and see if the stopper and the cage will fit. Then remove

the stopper and take the eudiometer to the hood.

4. **While wearing gloves**, measure out 20 mL of HCl using a graduated cylinder. Use a disposable pipette, to dispense 20 mL (±10%) of the HCl solution in the eudiometer. Next, carefully add the measured HCl into the eudiometer tube utilizing the hood ventilation.

5. Tilt the eudiometer slightly to the side (30-35 ° angle). Slowly and with caution, add deionized water to the eudiometer **WITHOUT DISTURBING** the acid that is already in the tube. Because of the different densities, you may observe two different layers. Add the water until the eudiometer is completely full, again avoiding mixing the DI water with the HCl layer. (Figure IG.2b).

6. Insert the stopper with the wire cage. No reaction should be taking place yet.

7. Place a gloved finger over the hole in the stopper and invert the eudiometer placing the stoppered end into a 400-mL or 600-mL beaker containing water. The water should be enough to cover the stopper (approx. 200 mL). Do not press the stopper onto the bottom of the beaker because liquid must be allowed to exit the hole in the stopper as the gas forms in the eudiometer tube.

8. Support the eudiometer with a buret clamp during the reaction, as shown in Figure IG.3. The acid should now slowly diffuse down though the inverted tube, and as the acid reaches the metal, bubbles will appear. These bubbles will rise to the top of the tube and collect. Allow the reaction to proceed until all the magnesium is gone and nothing remains but the copper wire cage.

Figure IG.3: Support the eudiometer using a buret clamp. (Note that while the rubber stopper is not shown in the image, it should be present in your setup.)

9. If pieces of magnesium metal remain or perhaps escape, then the experiment needs to be repeated. If you have produced more hydrogen gas than can be quantified by the eudiometer or if some of the gas escaped, then you need to repeat the experiment.

   At this point you might note that the volume of solution in the beaker is lower than the volume of solution in the eudiometer which means that the pressure inside the eudiometer is different than the pressure outside the eudiometer. Since we choose to use the barometric pressure (that outside the tube) as a measure of the pressure of the gas inside the tube, we need to get those two pressures equal, which will occur once the level of liquid inside the tube is equal to the level of liquid outside the tube. That is the purpose of the **“Equalization Chamber.”**

10. When the reaction is completed, with a finger remaining over the hole in the stopper, take the eudiometer to the tall container of water provided (the **Equalization Chamber**). Submerge the eudiometer (at which time you may remove your finger from the stopper) in the **Equalization Chamber** until the two levels of liquid (outside the eudiometer and inside the eudiometer) are at the same point and then read the volume of gas collected in the eudiometer while the eudiometer is still in the **Equalization Chamber**. At this time you will be measuring the volume when the pressure inside the tube is equal to the barometric pressure/atmospheric pressure.

11. You may now discard the liquid in the eudiometer tube in the appropriate waste container.

12. Your instructor or TA will show you how to read the barometer so that you can get a value of barometric pressure to the nearest 0.1 mmHg (millimeters of mercury). You will use the exact conversion definition, 760 mm Hg = 1 atmosphere, to convert from mm Hg to atmosphere. Record the air temperature in the laboratory.

13. Repeat the experiment. You may assume that the barometric pressure and the temperature are constant during the two trials.

Name: ______________________________

Section: __________ Date: ______________

Report Sheet:

Ideal Gas Law Experiment

Data and Calculations

Measured Data

	Trial 1	Trial 2
Mass of Mg	____________	____________
Volume of HCl (mL)	____________	____________
Volume of H_2 gas	____________	____________
Barometric pressure (mmHg)	____________	____________
Temperature (°C)	____________	____________

Calculations

	Trial 1	Trial 2
Moles of Mg reacted	____________	____________

Show calculations:

	Trial 1	Trial 2
Moles of H_2 formed	____________	____________

Show calculations (NOT calculated from molarity of HCl):

	Trial 1	Trial 2
Vapor pressure of water from Figure IG.1 (mmHg)	____________	____________
Pressure of H_2 from Dalton's law (mmHg)	____________	____________

Show calculations:

Pressure of H_2 in atm ____________ ____________

Show calculations:

Temperature in Kelvin ____________ ____________

Show calculations:

Calculation of R ____________ ____________

Show all work including any unit conversions:

Average value of R ____________

Show calculations:

Discussion

1. In your own words and complete sentences, explain the purpose of the equalization chamber.

2. In your own words and complete sentences, explain why the water vapor pressure was subtracted from P_{bar}? Where did the water vapor pressure come from in the experiment?

3. Describe in your own words and in complete sentences two reasons for your percent error.