

Experiment 13 - Calculation of the Ideal Gas Constant

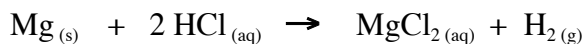
According to both theory and experiment, the pressure (P) of any sample of an ideal gas is inversely proportional to its volume (V) and directly proportional to its absolute temperature (T) and to the number of moles (n) of the gas present in the sample.

This proportionality may be written: $P \propto \frac{nT}{V}$ in which \propto is read "is proportional to." It

is more convenient to put this relationship in the form of an equation: $P = \frac{nRT}{V}$ in

which R, the "proportionality constant" is usually called the "ideal gas constant." This equation is more familiar in the rearranged form $PV = nRT$, and this equation is called the "ideal gas equation" or the "ideal gas law."

The object of the present experiment is to verify this equation for a sample of hydrogen gas, $H_{2(g)}$. To do this, we must measure P, V, T, and n. The value of R can then be calculated and compared with the accepted value of R. In this experiment, P, V, and T will be measured directly, while n will be calculated from the amount of magnesium used to produce the hydrogen gas according to the equation:



Safety Precautions:

- Wear your safety goggles.
- Use caution when handling 6 M HCl. It is corrosive and can burn your skin. If any HCl comes into contact with your skin, rinse it off immediately and thoroughly with lots of water.

Waste Disposal:

- At the end of the experiment, the HCl solution will be much more dilute. The water/HCl/ $MgCl_2$ mixture may be rinsed down the sink with plenty of water.

Procedure

1. Obtain a piece of magnesium ribbon approximately 5 cm long. Clean the ribbon with fine steel wool and weigh it accurately (to the nearest 0.0001 g). Make sure that it weighs less than 0.04 grams. (If it weighs more, cut off a small piece and re-weigh.)
2. Roll the ribbon loosely and then wrap it in a little ball of fine copper wire (see the display in the laboratory), leaving a "handle" of copper wire. The wrapping is designed to prevent small pieces of magnesium from breaking off and escaping during the experiment. Make sure that the ball is not too large to fit into the gas measuring tube.

3. Set up a ring stand with a buret clamp in position to hold a 50-mL gas-measuring tube (eudiometer). Fill a 400-mL beaker about 2/3 full of tap water and place it near the ring stand.
4. Tilt the gas-measuring tube slightly and pour in about 10 mL of 6 M HCl. (Estimate volume using the marks on the tube, and don't worry about getting exactly 10 mL.) Then, with the tube still in the same tilted position, gently add some water from a wash bottle, being careful not to mix the water too much with the acid. Then, gently fill the tube to the top with water (pour from a beaker or wash bottle). While pouring, rinse down any acid that may have wet the sides of the tube. The object is to have acid at the bottom of the tube and water at the top. Try to avoid stirring up the acid layer at the bottom of the tube. Air bubbles that may cling to the insides of the tube can be dislodged by gentle tapping of the tube.
5. Holding the copper coil by the handle, insert the cage about 3 cm down into the tube. Hook the wire over the edge of the tube where it will be pinched by the rubber stopper and held in place. When you insert the stopper, don't put your finger over the hole in the stopper. Let the water overflow as you insert the stopper so that there are absolutely no air bubbles trapped in the tube.
6. Add some more water to the hole in the stopper so that it is completely filled with water. Cover the hole in the stopper with your finger and invert the tube in the beaker of water, so that the stoppered end is under water. Once the hole is under water, you can remove your finger; the water cannot now run out. Clamp the tube in place. The acid, being more dense than the water, will diffuse down through the water and soon reach the metal sample. When the reaction begins, you will see bubbles of hydrogen gas form. Check to see if there is a temperature change associated with this reaction.
7. After the bubbles stop forming, you know that the reaction is completed, but you should wait for a few minutes for the tube to come to room temperature and for bubbles that may be clinging to the sides of the tube to be dislodged. (Tap, if necessary.)
8. Read and record the volume of the hydrogen gas in the eudiometer to the nearest 0.1 mL. Without changing the position of the eudiometer, hold a ruler at the top surface of the water level inside the beaker, and measure the distance from the water level in the beaker to the water level inside the eudiometer. Record this height (in centimeters). This will be the "height of the liquid column" referred to in the calculations. Measure and record the temperature of the gas by holding a thermometer against the eudiometer where it contains gas. Record the barometric pressure (the instructor will read the barometer and write today's atmospheric pressure on the board).
9. Disassemble the apparatus. If your instructor so directs, slowly add some sodium carbonate (Na_2CO_3) or sodium bicarbonate (NaHCO_3) to the HCl solution in the beaker until there is no further fizzing (this step neutralizes the remaining acid). Dump the neutralized solution down the drain. If there is no sodium carbonate or bicarbonate available, dump the solution down the sink followed by lots of water. Rinse the eudiometer with water.
10. Repeat the entire procedure with a second sample of magnesium.

11. Do your calculations, and write your name and your average value of R on the board in the classroom. Write down the class results in your laboratory notebook.

Calculations

1. From the mass of magnesium used and the balanced equation, calculate the number of moles of hydrogen gas expected for each trial.
2. To calculate the partial pressure of hydrogen in the tube, there are two corrections to take into account. First, the gas is collected over water, so the vapor pressure of water must be subtracted from the total pressure. The second correction is due to the column of solution remaining in the tube at the end of the reaction. Because the level of solution inside the tube is higher than the water level in the beaker at the end of the reaction, the pressure of gas inside the tube is lower than the atmospheric pressure. You will need to convert the height of the acid column to the corresponding height of mercury. Since mm Hg is a unit of pressure, it may then be subtracted from the total pressure.

To convert the height of the column of solution to the corresponding height of mercury, the differing densities of mercury ($d = 13.6 \text{ g/mL}$) and the acid solution ($d = 1.05 \text{ g/mL}$) must be taken into account. The equation $P = gdh$ relates the pressure due to a column of liquid to its height and density. (P is pressure due to a column of liquid, g is the constant acceleration of gravity, d is the density of the liquid, and h is the height of the column of liquid.) In a more useful form, you can relate the height of any liquid to the height of mercury that corresponds to the same pressure: $d_{\text{liq}}h_{\text{liq}} = d_{\text{Hg}}h_{\text{Hg}}$ (the constant g cancels out). This becomes:

$$h_{\text{Hg}} = \frac{h_{\text{liq}}d_{\text{liq}}}{d_{\text{Hg}}}, \text{ and this is the form of the equation you will use in your}$$

calculation. Make sure to convert the height of the liquid column to units of millimeters before using the above equation.

In summary:
$$P_{\text{atmosphere}} = P_{\text{H}_2} + P_{\text{water vapor}} + P_{\text{correction}}$$

To calculate the partial pressure of hydrogen, subtract the vapor pressure of water and the pressure correction (we'll abbreviate it as P_{corr}) due to the column of water from the total pressure. Do this for each trial.

3. Using the Ideal Gas Law, $PV = nRT$, calculate an experimental value for R for each trial.
4. Report the value of R obtained for each trial and the average value. Calculate the percent error using the theoretical value of R ($0.08206 \text{ L}\cdot\text{atm/K}\cdot\text{mol}$) and your average experimental value of R.

5. Using student results that have been listed on the chalkboard, calculate the standard deviation in the value of R for the class. You may either:
- Calculate the standard deviation, showing all of your work, for 5 student results, or
 - Calculate the standard deviation for the results of the entire class on your calculator.

Questions

- Why don't you need to measure the exact volume of acid used?
- Discuss any assumptions that we are making in the calculations for this experiment.
- In a similar experiment, a piece of aluminum was reacted with HCl. The hydrogen gas produced was collected in a eudiometer in the same way.
 - Write and balance the equation for this reaction. (Hint: what is the formula of aluminum chloride, the other product?)
 - If 39.5 mL of H₂ are produced at 21°C when the atmospheric pressure is 762.8 mmHg, and if the height of the liquid column in the eudiometer is 11.2 cm, what mass of aluminum was used?
- Once you have calculated the standard deviation of the student results for the class, explain its significance. (We have 95% confidence that the true value of R lies between _____ and _____, assuming only random errors.)