Gas Laws

Introduction: Although we cannot see gases, we can observe their behavior and study their properties. For example, we can watch a balloon filled with helium gas floating in air and conclude that helium gas is less dense than air. Similarly, a hot air balloon can ascend into the sky because the hot air in the balloon is less dense than the air around the balloon.

In addition, some common chemical reactions involve gases. For example, some antacids (e.g. Tums and Alka-Seltzer) contain compounds (e.g. $CaCO_3$ or $NaHCO_3$) that neutralize stomach acid—which is hydrochloric acid, HCl(aq)—to produce water and carbon dioxide gas. These reactions are shown below:

$$2\ HCl(aq) + CaCO_3(s) \rightarrow H_2O(l) + CO_2(g) + CaCl_2(aq)$$
$$HCl(aq) + NaHCO_3(s) \rightarrow H_2O(l) + CO_2(g) + NaCl(aq)$$

Furthermore, the airbags in our cars take advantage of another important gas reaction involving sodium azide powder, $NaN_3(s)$. When this powder is ignited, the substance quickly decomposes into sodium metal and nitrogen gas.

$$2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$$

The nitrogen gas inflates the airbags which generally have a total volume of 60.0 L. Because auto manufacturers need to make sure that the decomposition of the powder will not produce too much or too little gas, they use stoichiometry to determine the correct mass of powder to inflate the airbags. Thus, gas phase stoichiometry can be critical in determining the amount of starting material to use so that the appropriate volume of gas is generated in a reaction.

This lab will apply several concepts from Ideal Gas Laws. You will use your knowledge of chemical reactions and gas phase stoichiometry to **predict** the theoretical number of moles of gas formed as a product of the reaction of aluminum with hydrochloric acid. You will then use the Ideal Gas Law to predict the theoretical volume of the gas formed. Then you will carry out the actual reaction, measure the actual volume of gas generated, and use Dalton's law of partial pressures to calculate the actual number of moles of gas produced which will be compared to the theoretical moles predicted.

Example:

Let's pretend you swallowed a very small, 0.0145 gram, piece of magnesium metal. Your stomach contains hydrochloric acid for digestive purposes, which will react with the metal. If you could burp all the gas produced in one giant belch, what would the volume of the gas be at standard temperature and pressure (STP)*? (Notice we are predicting the volume by a theoretical calculation. You should not actually try this!)*

First we need a balanced chemical equation:

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

Using the mass of magnesium present calculate the moles of magnesium that would react and then use the mole ratio to find the moles of hydrogen gas produced:

$$0.0145 \text{ g Mg} \left(\frac{\text{mol}}{24.31\,\text{g}}\right)\left(\frac{1\,\text{mol H}_2}{1\,\text{mol Mg}}\right) = 5.96 \times 10^{-4} \text{ mol } H_2 \text{ gas}$$

Now use the Ideal Gas Law (PV=nRT) to calculate the volume of the hydrogen gas

$$V = \left(\frac{nRT}{P}\right) = \left(\frac{(5.96 \times 10^{-4}\text{ mol})\left(0.08206\,\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273\,\text{K})}{1.00\,\text{atm}}\right)\left(\frac{1000\text{ mL}}{\text{L}}\right) = 13.4 \text{ mL}$$

Note that the STP values are not applicable here, because the temperature of your stomach would not be 273K. We would also need to consider water vapor in addition to the hydrogen gas. Water in our stomachs evaporates and this water vapor would add volume to the belch.

When this type of reaction is carried out using equipment in the lab, the system can be closed so that no gas would escape (unlike your stomach, which has sphincters!). The pressure of a gas system can be controlled using a leveling bulb, which keeps the gas pressure equal to the pressure of the atmosphere (see the notes in the Experimental Set-up). When the level of water in the leveling bulb equals the level of water in an attached container like a buret, the pressure of the gas in the container is equal to atmospheric pressure.

Equipment Set-up and Leveling bulb discussion:

A buret will be used to collect our gas product, so we can measure the volume of gas produced accurately. A **leveling bulb** (a plastic bottle with an opening at its base) is attached to the bottom of the buret with a rubber hose. The opening allows the water in the system to be exposed to the atmospheric pressure in the lab. The leveling bulb is filled with water, and the water moves through the hose to fill the buret. As gas is produced, the water will be pushed out of the buret into the leveling bulb thus preventing excess pressure build up.

The **reaction chamber** is a round-bottomed flask, into which the hydrochloric acid and aluminum metal is placed. There is a second hose that connects the reaction vessel to the top of the buret. When the reaction flask is screwed on, the system is closed. **Note that when the reaction flask is removed by unscrewing the screw cap, the water levels in the buret and leveling bulb will be equal (both at atmospheric pressure)**. Thus, to raise or lower the water level in the buret you simply raise or lower the leveling bulb. Before beginning the reaction with the reaction flask off, adjust the leveling bulb up and down. Verify that as the bulb is raised, the water in the buret rises to match the level of water in the bulb and when the bulb is lowered, the water level in the buret drops to match the level in the bulb, so the water levels remain equal.

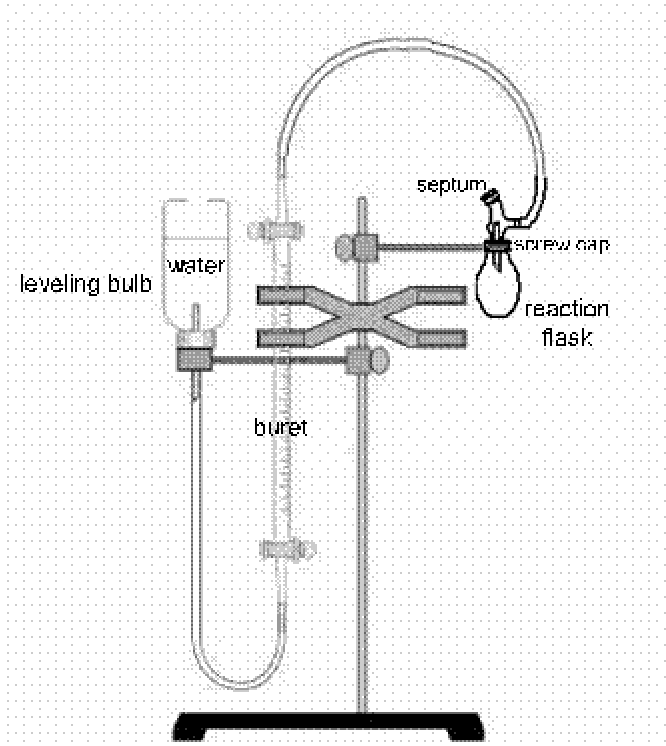


Figure 1: Experimental Set-up with Leveling Bulb

Check for leaks. With the reaction flask screwed on and the rubber septum attached, raise and lower the bulb. The water should not flow as freely as before due to the system being closed, and the water levels should not remain the same. With the reaction flask on (system closed), lower the bulb so the water level in the buret is higher than the water level in the bulb. The level of water in the buret should not continue dropping to the same level as the water level in the bulb. If it does, your system has a leak. Check to make sure the reaction flask is screwed on securely and the septum is attached. If you cannot find the source of the leak, consult your instructor.
Once leaks have been eliminated, lower the leveling bulb, so its water level is lower than the level of water in the buret. The pressure inside the buret is now less than atmospheric pressure (the gas in the buret is not pressing down as hard as the atmosphere). Now raise the bulb so that the water level in the buret is lower than the water level in the bulb. The pressure inside the buret is now higher than atmospheric pressure (the gas in the buret is pressing down harder than the atmosphere).

Remember, when the water levels are equal, the pressures are equal. **You will keep the water levels equal while running the reaction so that the pressure inside the reaction chamber and buret is equal to atmospheric pressure.**

Water Vapor Pressure Equations:

As a gas is collected over water it becomes saturated with water vapor. The total pressure in the system at the end of the reaction is, thus, due to both the gas product's pressure and the water vapor pressure. So the total pressure is the sum of the pressure of the gas (P_{gas}) and the pressure of the water vapor (P_{H_2O}). This is a result of Dalton's Law of Partial Pressures.

$$\boldsymbol{P_{total} = P_{gas}} + P_{H_2O}$$

When the water level in the buret is kept level with the water in the leveling bulb, the total pressure inside the buret is the same as the pressure *outside* the buret. Thus the total pressure is simply the atmospheric pressure. The pressure of the water vapor depends only on temperature and can be obtained by entering the temperature (in °C) into the Properties of water Website or by looking it up in the CRC Handbook of chemistry and physics.

(http://antoine.frostburg.edu/chem/senese/javascript/water-properties.html).

Thus, the pressure of the gas alone can be calculated by subtracting the vapor pressure of water from the total atmospheric pressure.

In reality, if a reaction is carried out in a closed container to trap any gas product, the pressure inside the container will build up. If too much pressure results, the container could explode. We must avoid building up pressure in the glass buret, and this is accomplished by keeping one end (the leveling bulb end) of the apparatus open to the atmosphere throughout the experiment. Thus, when gas is generated, the gas can push the water through the buret into the leveling bulb.

SAFETY PRECAUTIONS: *In the presence of a flame, the hydrogen gas given off in this reaction can react explosively with oxygen in the air. No flames are thus allowed in laboratory during this experiment.*

Hydrochloric acid *can cause burns if in comes in contact with skin. If any acid is spilled on your skin, wash immediately with soap and water, tell your instructor, and rinse for 15 minutes. Safety goggles must be worn at all times because contact with HCl(aq) for even a short time can cause permanent eye damage.*

Lab Notebook

Prepare a data table that contains the following information for 3 trials:

- Mass of aluminum
- Theoretical moles of gas produced (calculated before reaction)
- Theoretical volume of gas produced (calculated before reaction)
- Initial buret volume
- Final buret volme
- Volume of gas produced
- Actual pressure of gas (same for all 3 trials)
- Actual moles of gas produced
- Percent Yield

You will also need to record: room temperature, atmospheric pressure, and vapor pressure of water at room temperature for use in calculations.

Procedure:

1. Obtain 6 round pieces of an aluminum can by using a hole punch. Obtain three weigh boats, and put two of the aluminum pieces in each. *(Each trial will use 2 pieces.)* Weigh the aluminum pieces on an analytical balance by taring the empty weigh boat first. Record the **exact mass** of the aluminum pieces for your three trials in your data table.

2. Now carry out the following theoretical calculations in your LAB NOTEBOOK. (*Make sure to label each calculation such that it can be graded*):
   a. Write and balance the equation for the reaction between solid aluminum and hydrochloric acid.

   b. Use this balanced equation and stoichiometry to calculate the number of moles of gas that should be produced in the reaction based on the mass of aluminum. (*This is your theoretical number of moles of gas produced*.)

   c. Use the Ideal Gas Law to calculate the volume of gas that should form given the theoretical number of moles of gas calculated. (*This is your theoretical volume of gas produced*.)

      <u>Note:</u> You will use the atmospheric pressure (uncorrected for the vapor pressure of water) in this theoretical calculation. Since the gas *will* be collected over water, it serves as an *approximation* of how much gas will be collected during each trail in the experiment. Thus, we will not be comparing volumes to calculate percent yield.

      Have your instructor confirm your calculations and initial in your lab notebook.

   d. Do NOT forget to put your results from calculations b. and c. above in your data table.

3. Now you are ready to test your theoretical calculations. Make sure your equipment is set up according to Figure 1, shown on page 1 of this handout.

   a. Adjust the leveling bulb and check for leaks as described in the Experimental Set-up section.

   b. Take your graduated cylinder and obtain about 3.0 mL of 6M hydrochloric acid at the reagent station. Return to your bench and take care not to spill any of the acid.

   c. Remove the reaction flask, and pour the hydrochloric acid inside.*(The flask does not need to be dry – if you are unsure if it is clean rinse with DI water several times.)*

   d. Lift the leveling bulb so that the buret water level is around 15 mL. When everyone is ready in your group, drop the two aluminum pieces for trial one into the flask. Screw the reaction flask back on while holding the leveling bulb steady so that the water levels in the bulb and buret are equal. Quickly record the exact volume in the buret as the initial volume for trial 1 making sure the bulb and buret water levels are still equal (the reaction takes a couple of seconds to get started…so *quickly* record the volume.).

   e. While the reaction proceeds, keep the leveling bulb and buret water levels equal.

   f. When the reaction has stopped, hold the bulb steady with the water levels equal, and record the exact final buret volume.

   g. Unscrew the reaction flask, empty the contents into your waste beaker, and rinse with DI water twice.

4. Repeat step 3 for the remaining two trials with the other aluminum can pieces.

5. When you are done with all 3 trials, wash the reaction flask using tap water and then rinse the reaction flask with DI water for the next class. **Do the same for all your glassware and wipe down your lab bench.**

6. ***LAB NOTEBOOK:*** Using your results explicitly label and complete the following calculations in your lab notebook. Do not forget to complete your data table.

   a. Pressure due to the gas produced in reaction. To calculate this, you will the atmospheric pressure and subtract the vapor pressure of water (See the Water Vapor Pressure section, p. 2).

   b. Actual moles of gas produced for each trial. In this calculation, you will use the pressure that you collected in the previous step.

   c. Percent yield for each trial.

   d. Average percent yield.

7. ***LAB REPORT:*** Submit the following for your lab report: Lab Notebook copies and Postlab Questions (pages 7-8).

Gas Laws: *Lab Report*

Name: ________________________
Partner(s): _____________________
Section Number: _______________

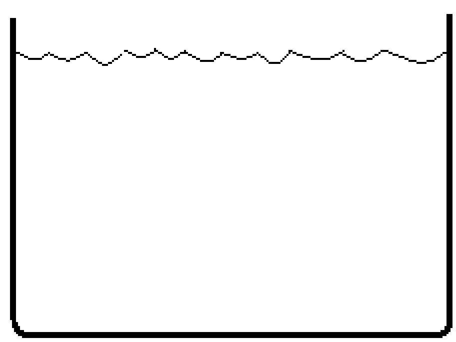
Post Lab Questions:

1. Using the Activity Series circle the metals that would react with hydrochloric acid to produce hydrogen gas.

   Cu(s) Ni(s) Au(s) Mg(s) Pb(s) Ag(s) Mn(s) Fe(s)

2. Why was the exact volume of the hydrochloric acid in the flask not recorded?

3. Draw a picture representing 5 molecules of hydrochloric acid in the beaker of water. *Hint: Is hydrochloric acid a strong, weak, or non-electrolyte?*



4. Would the following 3 hypothetical errors cause a higher or lower percent yield? Explain your answer in each case.

   a. A leak in the rubber tubing allows some gas to escape.

High Low

   b. A student forgot to subtract the vapor pressure of water from the atmospheric pressure while doing the calculations.

High Low

c. A student held the leveling bulb way below the buret's water level while reading the final buret volume.

High **Low**

5. The reaction between the hydrochloric acid and the can began slowly. When the reaction began to accelerate, two thin films of plastic came off the metal. These plastic films did not react with the acid. Why did the reaction start slowly and then suddenly accelerate?

6. There is a difference between your theoretical and actual volume of hydrogen gas. Assuming your technique was good and there was no human error, explain the difference.

7. In reality, most of the gas in the buret was actually air from the tubing and reaction flask that was pushed into the buret as hydrogen gas was produced. Does the presence of this air (mixture of gases) instead of hydrogen gas affect your experiment? Explain your answer.