MOLAR VOLUME OF A GAS

Background: In many cases, the amount of gas evolved by a reaction is of interest. Since gases have such small densities, it is usually not practical to collect the gas and find its mass. For gases that are not particularly soluble in water, it is possible to collect the evolved gas by displacement of water from a container.

The setup for the collection of a gas over water involves a container in which the reaction takes place and a gas collection container filled with water and inverted in a reservoir of water. The gas evolved from the reaction is collected by attaching one end of a hose to the reaction container and inserting the other up into the inverted gas collection bottle. As the gas is created, it will displace water from the bottle. The volume of gas can be determined by the amount of water that was displaced by the gas.

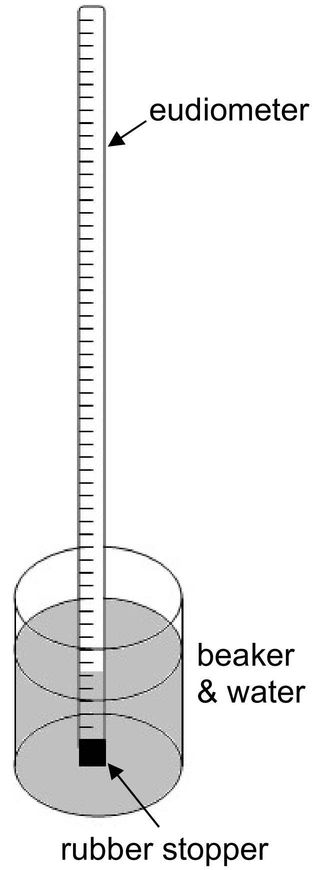

Figure 1

The volume of gas collected and the gas laws can be used to calculate the number of moles of gas collected.

During the collection, the water level in the container will adjust so that the pressure inside and outside the container are the same. Because of this, if we know the atmospheric pressure, we also know the pressure of the gas inside the bottle.

The pressure inside the bottle is partially from the gas being collected and partially from the water vapor that has escaped from the surface of the water in the jar. The water inside the jar will reach an equilibrium state where the number of molecules leaving the surface is the same as the number returning. The equilibrium pressure of water is temperature dependent and is called the **vapor pressure of water.**

Dalton's Law of Partial Pressures tells us that the total pressure in the container must be the sum of the pressures of the gas we collected and the water vapor. $\mathbf{P_T = P_{gas} + P_{H2O}}$

This equation can be used to calculate the pressure of the gas collected. Once the pressure of the collected gas is known, the number of moles of gas can be calculated using the ideal gas law.

In today's experiment, you will calculate a molar volume for hydrogen gas at STP as well as calculating your own value for the universal gas constant, R.

Key Concepts:

Avogadro's Law	Dalton's Law	Ideal gas law	Molar volume

Pre-lab Questions:

1. Write the balanced chemical equation for the reaction of solid magnesium and aqueous hydrochloric acid. What type of reaction is this?
2. Use Dalton's Law of Partial Pressures to explain why the pressure of the hydrogen gas will be less than the observed barometric pressure in the lab.
3. A reaction of 0.028 g of magnesium with excess hydrochloric acid generated 31.0 mL of hydrogen gas. The gas was collected by water displacement in a 22°C water bath. The barometric pressure in the lab that day was 746 mm Hg.
    a. Calculate the partial pressure of hydrogen gas in the gas collecting tube.
    b. Use the combined gas law to calculate the corrected volume of hydrogen at STP.
    c. What is the theoretical number of moles of hydrogen that can be produced?
    d. Calculate the molar volume of hydrogen at STP.
4. Create a data table and a hazards table.
5. Draw a sketch of the procedure.

Materials:

400 mL beaker	balance or ruler	ring stand	copper wire (optional)
piece of Mg ribbon	eudiometer	clamp	thermometer
1 M HCl	rubber stopper	distilled water	

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Table 1: Vapor Pressure of Water at Different Temperatures

Temperature, °C	P_{H2O}, mm Hg	Temperature, °C	P_{H2O}, mm Hg	Temperature, °C	P_{H2O}, mm Hg
16	13.6	20	17.5	24	22.4
17	14.5	21	18.7	25	23.8
18	15.5	22	19.8	26	25.2
19	16.5	23	21.1	27	26.7

Procedure:

1. Fill a 400 mL beaker about half full of water. Stand to one side so that the temperature of the water can come to room temperature.
2. Obtain a piece of magnesium (Mg) ribbon no more than 2.0 cm long. Measure and record the length of your sample of ribbon to the nearest tenth of a mm. Fold the magnesium ribbon so that it will fit inside the end of the eudiometer (gas measuring tube).
3. Prepare a ring stand and clamp to support the eudiometer in the position shown in Figure 1.
4. Slowly pour 20 mL of 1 M HCl into the eudiometer.
5. Incline the tube slightly so that air may escape. Slowly fill the eudiometer with distilled water from a beaker. Take care to mix the acid and the water as little as possible. Pouring the water down the side of the tube slowly can help with this.
6. With the tube completely full of water, place the magnesium ribbon into the tube. Quickly stopper the eudiometer. The stopper should force water and any remaining air bubbles out of the tube.
7. With your finger over the hole in the stopper, invert the tube and place the stoppered end in the beaker of water. Clamp the eudiometer in place so that the bottom of the rubber stopper IS NOT TOUCHING the bottom of the beaker.
8. When the magnesium has reacted completely and the evolution of gas has stopped, tap the tube gently with your finger to dislodge any bubbles you see attached to the side of the tube.
9. Place your finger over the hole in the stopper and remove the tube from the beaker. Lower the tube into the larger container of water used to equalize pressure and remove your finger. Raise or lower the tube until the level of the water inside the eudiometer is the same as the level outside. Read the volume of the eudiometer as accurately as possible and record this volume.
10. Empty the contents of the eudiometer and the beaker and rinse both with tap water.
11. Record the room temperature and barometric pressure from the board.
12. Repeat procedure 1-2 times to obtain data for more than one trial.

Calculations:

1. Calculate the mass of magnesium in the sample used.
2. Calculate the moles of magnesium in the sample used.
3. Calculate the partial pressure (in atm) of hydrogen gas in the hydrogen gas/water vapor mixture.
4. You have the volume, pressure, and temperature of the gas. Calculate the volume that would be occupied at STP.
5. Calculate the molar volume (L/mol) of your sample at STP.
6. Calculate a percent error using 22.4 L/mol as the theoretical value.
7. Using your values for pressure, temperature, volume, and moles, calculate a value for the ideal gas constant, R.
8. Calculate a percent error for your calculated value of R using 0.0821 L*atm/mol*K as the theoretical value.

Questions to ponder...and answer!

1. One mole of hydrogen gas has a mass of 2.02 g. Use your value of the molar volume of hydrogen to calculate the mass of one liter of hydrogen gas at STP. This is the density of hydrogen in g/L. How does this experimental value of density compare with the literature value?
2. Why is it important for the stopper to have a hole in it? Why have a stopper at all?
3. Why was it important to make sure the eudiometer was filled completely with water before you inverted it?
4. A student noticed the magnesium ribbon appeared to be oxidized—the metal surface was black and dull rather than silver and shiny. What effect would this error have on the measured volume of hydrogen gas? Would the calculated molar volume of hydrogen be too high or too low as a result of this error? Explain.