Molar Volume of Hydrogen
Combining the Gas Laws

Introduction

Airbags have been required safety features on new cars since the 1980s and are credited with saving thousands of lives over that time. Airbags contain a compound that decomposes to give nitrogen gas upon impact from a collision. How much nitrogen gas must be generated to fill an airbag? The amount of gas needed to fill any size container can be calculated if we know the molar volume of the gas.

Concepts

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

Background

Avogadro's law states that equal volumes of gases contain equal numbers of molecules under the same conditions of temperature and pressure. It follows, therefore, that all gas samples containing the same number of molecules will occupy the same volume if the temperature and pressure are kept constant. The volume occupied by one mole of a gas is called the *molar volume*. In this experiment we will measure the molar volume of hydrogen gas at standard temperature and pressure (STP, equal to 273 K and 1 atm).

The reaction of magnesium metal with hydrochloric acid (Equation 1) provides a convenient means of generating small-scale quantities of hydrogen in the lab.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g) \qquad \textit{Equation 1}$$

If the reaction is carried out with excess hydrochloric acid, the volume of hydrogen gas obtained will depend on the number of moles of magnesium as well as on the pressure and temperature. The molar volume of hydrogen can be calculated if we measure the volume occupied by a sample containing a known number of moles of hydrogen. Since the volume will be measured under laboratory conditions of temperature and pressure, the measured volume must be corrected to STP conditions before calculating the molar volume.

The relationship among the four gas variables—pressure (P), volume (V), temperature (T), and the number of moles (n)—is expressed in the ideal gas law (Equation 2), where R is a constant called the universal gas constant.

$$PV = nRT \qquad \textit{Equation 2}$$

The ideal gas law reduces to Equation 3, the combined gas law, if the number of moles of gas is constant. The combined gas law can be used to calculate the volume (V_2) of a gas at STP (T_2 and P_2) from the volume (V_1) measured under any other set of laboratory conditions (T_1 and P_1). In using either the ideal gas law or the combined gas law, remember that temperature must be always be expressed in units of kelvins (K) on the absolute temperature scale.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$ *Equation 3*

Hydrogen gas will be collected by the displacement of water in an inverted graduated cylinder using the apparatus shown in Figure 1. The total pressure of the gas in the cylinder will be equal to the barometric (air) pressure. However, the gas in the cylinder will not be pure hydrogen. The gas will also contain water vapor due to the evaporation of the water molecules over which it is being collected. According to Dalton's law, the total pressure of the gas will be equal to the partial pressure of hydrogen plus the partial pressure of water vapor (Equation 4). The vapor pressure of water depends only on the temperature (see Table 1).

$$P_{total} = P_{H_2} + P_{H_2O}$$ *Equation 4*

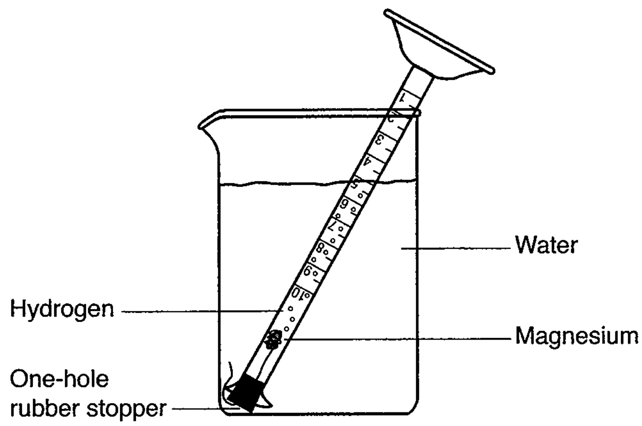

Figure 1.

Table 1. Vapor Pressure of Water at Different Temperatures

Temperature, °C	P_{H_2O}, mm Hg	Temperature, °C	P_{H_2O}, mm Hg
16 °C	13.6	22 °C	19.8
17 °C	14.5	23 °C	21.1
18 °C	15.5	24 °C	22.4
19 °C	16.5	25 °C	23.8
20 °C	17.5	26 °C	25.2
21 °C	18.7	27 °C	26.7

Experiment Overview

The purpose of this experiment is to determine the volume of one mole of hydrogen gas at standard temperature and pressure (STP). Hydrogen will be generated by the reaction of a known mass of magnesium with excess hydrochloric acid in an inverted graduated cylinder filled with water. The volume of hydrogen collected by water displacement will be measured and corrected for differences in temperature and pressure in order to calculate the molar volume of hydrogen at STP.

Pre-Lab Questions

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 water bath 22 °C. The barometric pressure in the lab that day was 746 mm Hg.

1. Use Dalton's law and the vapor pressure of water at 22 °C (Table 1) to calculate the partial pressure of hydrogen gas in the gas collecting tube.
2. Use the combined gas law to calculate the "corrected" volume of hydrogen at STP. *Hint:* Watch your units for temperature and pressure!
3. What is the theoretical number of moles of hydrogen that can be produced from 0.028 g of Mg? *Hint:* Refer to Equation 1 for the balanced equation for the reaction.
4. Divide the corrected volume of hydrogen by the theoretical number of moles of hydrogen to calculate the molar volume (in 4 mol) of hydrogen at STP.

Materials

Copper wire, Cu, 18-gauge, 10-cm long	Beaker, 400-mL
Hydrochloric acid, HCl, 2 M, 10 mL	Graduated cylinder, 10-mL
Magnesium ribbon, Mg, 1-cm pieces, 2	Metric ruler
Distilled or deionized water	One-hole rubber stopper, size 1 or 2
Wash bottle	Scissors or wire cutter
Barometer	Thermometer

Safety Precautions

Hydrochloric acid is a corrosive liquid. Avoid contact with eyes and skin and clean up all spills immediately. Magnesium metal is a flammable solid. Keep away from flames and other sources of ignition. Wear chemical splash goggles and chemical-resistant gloves and apron. Wash hands thoroughly with soap and water before leaving the laboratory.

Procedure

1. Fill a 400-mL beaker about ¾-full with water.
2. Obtain or cut a 1-cm piece of magnesium ribbon. Measure and record the exact length of the magnesium ribbon to the nearest 0.1 cm. Do not exceed 1.0 cm.
3. Your teacher will provide a conversion factor in g/cm to calculate the mass of magnesium used in this experiment. Multiply the length of magnesium ribbon by this conversion factor to calculate the mass of the 1-cm piece of magnesium obtained in step 2. Record the mass of magnesium in the data table.
4. Obtain a piece of copper wire about 10-cm long. Twist and fold one end of the copper wire around a pencil to make a small "cage" into which the magnesium ribbon may be inserted. See Figure 2.

Figure 2.

5. Firmly place the 1-cm piece of magnesium into the copper-wire cage.

6. Insert the straight end of the copper wire into a one-hole rubber stopper so that the cage end containing the magnesium is about 1-cm below the bottom of the stopper (see Figure 1). Hook the end of the copper wire around the top of the stopper to hold the cage in place.

7. Obtain about 5 mL of 2 M hydrochloric acid in a tall-form, 10-mL graduated cylinder.

8. While holding the graduated cylinder in a tipped position, slowly and carefully fill the graduated cylinder with distilled water from a wash bottle or a plastic pipet. Work slowly to avoid mixing the acid and water layers at this time. Fill the graduated cylinder all the way to the top so that no air remains in the cylinder.

9. Insert the magnesium–copper wire–stopper assembly into the graduated cylinder. The magnesium piece should be above the 10-mL line on the graduated cylinder (see Figure 1).

10. Place your finger over the hole of the rubber stopper, invert the graduated cylinder, and carefully lower the stoppered end of the graduated cylinder into the 400-mL beaker containing water.

11. Record any evidence of a chemical reaction in the data table.

12. If the magnesium metal "escapes" its copper cage, gently shake the graduated cylinder up and down to work it back into the acidic solution.

13. Allow the apparatus to stand for 5 minutes after the magnesium has completely reacted. Gently tap the sides of the graduated cylinder to dislodge any gas bubbles that may have become attached to the sides.

14. Gently move the graduated cylinder up or down in the water bath until the water level inside the graduated cylinder is the same as the water level in the beaker. This is done to equalize the pressure with the surrounding air (barometric pressure). *Note:* Be careful to make sure that the stoppered end of the graduated cylinder remains submerged in the water.

15. When the water levels inside and outside the cylinder are the same, measure and record the exact volume of hydrogen gas in the graduated cylinder.

16. Since the cylinder is upside down in the water bath, the volume reading obtained in step 14 must be corrected for the fact that the meniscus was read upside down as well. Subtract 0.2 mL from the volume recorded in step 15 and record the "corrected" volume of hydrogen gas in the data table.

17. Measure and record the temperature of the water bath in the beaker. Using a barometer, measure and record the barometric pressure in the lab.

18. Remove the graduated cylinder from the water bath and discard the water in the beaker as directed by your instructor.

19. Repeat the entire procedure to obtain a second set of data. Record this as Trial 2 in the data table. If time permits, perform a third trial as well.

Name: ______________________________

Class/Lab Period: ______________________

Molar Volume of Hydrogen

Data Table

	Trial 1	Trial 2
Length of Mg Ribbon		
Mass of Mg		
Evidence of Chemical Reaction		
Volume of H_2 Gas		
Corrected Volume of H_2		
Temperature of Water Bath		
Barometric Pressure		

Post-Lab Calculations and Analysis

Construct a Results Table to summarize the results of the following calculations.

1. Calculate the theoretical number of moles of hydrogen gas produced in Trials 1 and 2.

2. Use Table 1 in the *Background* section to find the vapor pressure of water at the temperature of the water bath in this experiment. Calculate the partial pressure of hydrogen gas produced in Trials 1 and 2.

3. Use the combined gas law to convert the measured volume of hydrogen to the volume the gas would occupy at STP for Trials 1 and 2. *Hint:* Remember the units!

4. Divide the volume of hydrogen gas at STP by the theoretical number of moles of hydrogen to calculate the molar volume of hydrogen for Trials 1 and 2.

5. What is the average value of the molar volume of hydrogen? Look up the literature value of the molar volume of a gas in your textbook and calculate the percent error in your experimental determination of the molar volume of hydrogen.

$$\text{Percent error} = \frac{|\text{Experimental value} - \text{Literature value}|}{\text{Literature value}} \times 100\%$$

6. 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 the density compare with the literature value? (Consult a chemistry handbook for the density of hydrogen.)

7. In setting up this experiment, a student noticed that a bubble of air leaked into the graduated cylinder when it was inverted in the water bath. What effect would this 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.

8. A student noticed that 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.

9. *(Optional)* Your teacher wants to scale up this experiment for demonstration purposes and would like to collect the gas in an inverted 50-mL buret. Use the ideal gas law to calculate the maximum length of magnesium ribbon that your teacher should use.