THE MOLAR RELATIONSHIP INVOLVING MASS AND VOLUME 28

Both solids and gases must often be handled in the same experiment. The amount of solid used or produced can be determined by measuring the mass of the material on a balance. It is difficult, however, to find the mass of a gas. For convenience the chemist measures gas volume and uses this value to calculate the mass of the gas. Therefore, it is necessary for the chemist to know the quantitative relationship between the molar mass and the molar volume of a gas. **Avogadro's hypothesis** explains the relationship between the molar volume, the molecular mass, and the actual mass of a sample gas. This hypothesis states that *equal volumes of gases under the same temperature and pressure contain equal numbers of molecules.* Also, the volume occupied by one mole of any gas at standard temperature and pressure equals 22 400 cm^3.

In this experiment you will investigate the chemical significance of Avogadro's hypothesis. You will determine the volume of hydrogen gas evolved in a reaction between magnesium metal and hydrochloric acid, and from your results determine the mass of H_2 produced. Your experimental results will then be compared to the results predicted by Avogadro's hypothesis. You will need to convert room temperature and pressure to standard conditions (STP) in order to compare your results.

Objectives

In this experiment, you will

- measure the mass of a piece of magnesium ribbon,
- react the magnesium with HCl(aq) and collect the gaseous product,
  measure the volume of the gas collected and convert the volume to standard conditions (STP), and
- calculate the molar relationship between the solid magnesium consumed and the gas produced.

EQUIPMENT

goggles and apron
beaker (250 or 400 cm^3)
thermometer
barometer
balance
Cu wire or cotton thread (15 cm)
ring stand
utility clamp
gas measuring tube (50 cm^3)
one hole rubber stopper (to fit tube)
battery jar (2 or 3 dm^3)

PROCEDURE

1. On entering the lab, fill a 250 or 400 cm^3 beaker about one-half full of tap water and allow it to stand so that the temperature of the water may adjust to room temperature.
2. Prepare a data table as directed in the Analysis. Record the temperature of the room, the barometric pressure, and the precise mass of 1.000 meter of magnesium ribbon (from your teacher). Safety goggles and lab apron must be worn for this experiment.
3. Cut a piece of magnesium ribbon, Mg, about 5 cm long. Make sure you cut the ends of the ribbon squarely. Carefully measure the length of your sample of ribbon to the nearest millimeter (0.1 cm). Record the length on your data sheet.
4. Obtain a piece of copper wire or cotton thread about 15 cm long and tie it to the magnesium ribbon which has been folded to a size that will fit inside the gas-measuring tube.
5. Prepare a ring stand with a utility clamp to support the gas-measuring tube.
6. Slowly pour about 10 cm^3 of 3*M* HCl into the gas tube. **CAUTION:** *Dilute HCl will stain, cause mild burns, and irritate lungs and eyes. Avoid contact and inhalation. Rinse spills with plenty of water.*
7. Incline the tube slightly so the air may escape and slowly fill it with tap water from the beaker. Pour the water slowly down the side of the tube so the water and acid mix as little as possible. Fill the tube completely. Refill the beaker three-fourths full of water.
8. With the tube completely full of water, insert the magnesium ribbon about 3 or 4 cm into the tube. With the wire or thread against the side of the

tube, insert a 1-hole stopper. The stopper should force water and all air bubbles out of the tube and should hold the thread or wire suspending the magnesium in place.

9. With your finger over the hole in the stopper (make sure there is no air in the hole of the stopper), invert stoppered end of the tube in the beaker of water. Clamp the tube in place so that the bottom of the rubber stopper is slightly above the bottom of the beaker. The reaction will not start immediately because it takes time for the acid to diffuse down through the column of water to the metal.
10. When the magnesium has reacted completely and evolution of gas has stopped, tap the tube with your finger to dislodge any bubbles you see attached to the side of the tube.
11. Place your finger over the hole in the stopper and remove the tube from the beaker. Lower the tube into a larger container of water (provided by your teacher) and remove your finger. Raise or lower the tube until the level of the water inside the tube is the same as the level of water outside the tube. This equalizes the pressure. Read the scale on the tube as accurately as possible (to the nearest 0.1 cm^3). This reading will give the volume of the gases (hydrogen and water vapor) in the tube. Record the volume.
12. Empty the contents of the tube and beaker into the sink, and rinse both with tap water.
13. Repeat the entire experiment if directed to do so by your teacher.

ANALYSIS

1. Prepare a table for your data. Use Table 28-1 as a guide. To convert temperature and pressure use K = C° + 273 and 1 kPa = 7.50 mm Hg.
2. Calculate the mass of magnesium by using the known mass per meter of ribbon. Assume the magnesium ribbon was of uniform thickness and width. (Hint: 5 cm = 0.05 meter.) Calculate the moles of magnesium in the sample.
3. Calculate the partial pressure of hydrogen gas in the hydrogen gas-water vapor mixture. See Table A-6 of the Appendix for the partial pressure of water vapor at different temperatures. Also, remember that the total pressure of a gas mixture is equal to the sum of the partial pressures of each gas:

$$P_{total} = P_{H_2O} + P_{H_2}$$

4. Use the gas laws to calculate the volume that would be occupied by the gas at standard temperature, 273 K, and standard pressure, 101.3 kPa. Use the volume of the hydrogen gas (from your experiment) at the corrected pressure (Calculation 3) and room temperature.
5. From your calculations, a fractional part of a mole of magnesium gave an experimentally determined volume of H_2 gas at STP. Use this information to calculate the volume of H_2 gas that could be produced if one mole of magnesium were reacted with excess HCl at STP.

Table 28-1

Mass of 1 meter Mg ribbon	.05	g
Length of Mg ribbon		cm
Moles of Mg ribbon used	.002	mol
Room temperature	22°	K
Barometric pressure	742 Torr	kPa
Volume of gas collected	50	cm^3
Corrected pressure of dry H_2 gas ($P_{H_2} = P_{room} - P_{H_2O}$)	722.2 Torr	kPa
Volume of H_2 gas at STP		cm^3
Volume of H_2 from 1 mole Mg at STP		cm^3

CONCLUSIONS

1. The reactants in this experiment are known. One of the products is hydrogen gas. If the water in the tube is evaporated, the other product, a white salt remains. Write the balanced chemical equation for this reaction. What is the molar relationship between the solid Mg and the H_2 gas?
2. Calculate the volume of gas you would expect to be evolved if *X* mol of magnesium were reacted with an excess of HCl at standard conditions.
3. From your experimental results, predict the volume occupied by one mole of hydrogen gas (molar volume) at STP. Compare this value to the theoretical value for the volume of one mole of H_2 at STP. Calculate your percentage error. What were your major sources of error?
4. If the experiment was conducted a second time, calculate the difference in your results. Explain the reasons for the difference in results.