

# 9 The Molar Volume of Hydrogen

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- Measure the mass and volume of a sample of hydrogen gas.
- Use the ideal gas law with these measurements to calculate the molar volume of hydrogen at standard temperature and pressure..

## DISCUSSION

For a specific amount of any ideal gas, the relation between the pressure of the gas,  $P$ , its volume,  $V$ , its temperature,  $T$ , and its number of moles,  $n$ , is given by the ideal gas law:

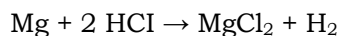
$$PV=nRT$$

Here  $R$  is the proportionality constant, called the gas constant, and it has the same value for all ideal gases under all conditions, namely,  $0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$ . In all calculations in which this constant is employed, pressure must be expressed in atmospheres, volume in liters, and temperature in Kelvins.

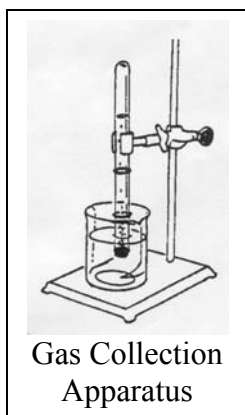
Standard conditions are defined as exactly 1 atm pressure and  $0^\circ\text{C}$  (273 K). The molar volume of a gas is the volume that 1.000 mol of it occupies under these conditions. It is the same for all ideal gases:

$$V = \frac{nRT}{P} = 22.414\text{L}$$

In this experiment, you will produce hydrogen gas,  $\text{H}_2$ , by reacting magnesium with hydrochloric acid according to the following equation:



You will assume the hydrogen is an ideal gas, and you will measure its mass, volume, temperature, and pressure. From these measured values, you will calculate the molar volume of hydrogen and compare the result with the ideal value above.



Gas Collection Apparatus

## PROCEDURE

1. Obtain a strip of magnesium ribbon. This strip has been pre-measured so that it will not produce more hydrogen than the collection tube will hold. Weigh the ribbon to the nearest milligram and record the mass on the report sheet.

Produce and collect the hydrogen gas as follows:

Claim a gas collection tube (already set up in the lab). Make sure it has two rubber bands on it, as shown in the diagram.

Fold up the magnesium ribbon into a small, tight bundle. Wrap it in all directions with copper wire so that the Mg is enclosed in a wire cage. It needs to stay confined even after the acid has eaten most of it away; therefore, be sure there are no significant openings in the cage. Leave 10 or 15 cm of wire extending from the cage.

Add about 10 mL of concentrated HCl to the gas collection tube. **CAUTION.** Be sure to use **hydrochloric acid**; others might react violently when the water is added. Then fill the tube completely with tap water.

Place the cage in the mouth of the tube so that it is about 3 cm below the surface of the water. Fold the wire extension over the side of the tube. Insert a one- or two-hole stopper into the opening so that the cage is held firmly in place.

Holding your finger over the stopper hole(s), invert the tube into a 400 mL beaker that is about half-filled with water. Then clamp the tube in place as shown in the diagram, with its mouth below the water's surface. (There is no need to rush this maneuver. The acid will take more than a minute to diffuse down to the stopper, and by then it

becomes dilute enough not to harm your finger.)

Observe the reaction. When no more hydrogen bubbles are visible, the reaction is complete. Wait an additional 5 minutes so that the hydrogen gas comes to room temperature. Move one rubber band to the level of the water in the beaker, and move the other to mark the water level in the tube.

2. Measure and record room temperature. Remove the tube from the beaker, and empty the liquid into the beaker.

3. Read the barometer in the corner of the lab, and record the atmospheric pressure in *torr* (*mm* of Hg).

4. Measure the distance between the rubber bands in millimeters. This is how high the water in the tube rose above the surface of the water in the beaker.

5. Fill the tube with tap water up to the first rubber band; *i.e.* the rubber band nearest the tube's closed end. This amount of water has the same volume as the hydrogen did. Pour it into a graduated cylinder, and record its volume.

6. Obtain the vapor pressure of water from the table.

7. Since mercury is 13.6 times as dense as water, a column of mercury will rise only 1/13.6 times as high as a column of water. Calculate the mercury equivalent of the water column by dividing its height by 13.6. The quotient comes out in *mm* of Hg or *torr*.

Temperature	Vapor Pressure (torr)
20	17.5
21	18.7
22	19.8
23	21.1
24	22.4
25	23.8
26	25.2
27	26.7
28	28.3
29	30.0

8. To find the partial pressure exerted by the hydrogen, you must recognize that the atmospheric pressure, measured in Step 3, equals the partial pressure of hydrogen gas in the tube, plus the partial pressure of water vapor mixed with the hydrogen, plus the pressure of the water column between the rubber bands (according to the equation below).

$$P_{atm} = P_{H_2} + P_{H_2O} + P_{water\ column}$$

Solve for the partial pressure of the hydrogen by subtracting the water's vapor pressure and the mercury equivalent of the water column from the atmospheric pressure.

9. Calculate the absolute (kelvin) temperature.

10. From the mass of magnesium calculate the moles of magnesium reacted.

11. Using the balanced equation, calculate the moles of hydrogen produced.

12. Using the combined gas law equation (shown below), convert your measured volume of hydrogen (from Step 5) to conditions of STP. Give the result in units of liters (*L*).

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

13. Calculate the molar volume by dividing the calculated volume at STP by the moles of hydrogen produced.

14. Compare your experimental molar volume to the theoretical value (22.414 *L*) by calculating the percent error in the molar volume.

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Name \_\_\_\_\_  
Partner \_\_\_\_\_  
Section \_\_\_\_\_ Locker \_\_\_\_\_  
Instructor \_\_\_\_\_

Enter the data or answer the questions *according to the corresponding step in the procedure*

1. Mass of magnesium strip	
2. Temperature of the hydrogen	
3. Barometric pressure	
4. Height of the water column	
5. Volume of hydrogen	
6. Vapor pressure of water	
7. Mercury equivalent of water column	
8. Pressure of hydrogen gas	
9. Absolute temperature	
10. Moles of magnesium reacted	
11. Moles of hydrogen produced	
12. Volume of hydrogen at STP	
13. Molar volume of hydrogen at STP (in <i>L</i> )	
14. Percent error	

## CALCULATIONS

**APPLICATION OF PRINCIPLES**

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1. Tell whether the following errors would **increase**, **decrease**, or have **no effect** on the experimental molar volume.

a) the measured mass of the magnesium was too small

b) the rubber band that marks the location of the ending water level in the tube was positioned too low

c) the actual temperature of the hydrogen is less than room temperature.

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2. All real gases deviate to some extent from the behavior of ideal gases. At standard conditions, the density of O<sub>2</sub> gas is 0.0014290 g/mL, that of H<sub>2</sub> gas is 0.00008988 g/mL, and that of CO<sub>2</sub> is 0.0019769 g/mL.

Using these values and the molecular weights from your textbook (in g/mol, not rounded off), calculate the molar volume of each of these, in mL/mol, to five significant figures. (Hint: analyze the units for density, molecular weight, and molar volume.)

Correlate the values of these three gases with the molar volume of an ideal gas (22414 mL/mol).

Which gas deviates least?

Which gas deviates most?

Suggest a reason for this behavior.