

Determination of the Universal Gas Constant, R

Objective: To investigate the relationship between the number of moles and the volume occupied by a gas at a given temperature and pressure; to use these data to estimate the value of the universal gas law constant, R

Materials: 1.0 M HCl; magnesium ribbon (Mg)

Equipment: Gas collection apparatus (pre-assembled; see Figure 1); ring stands and clamps; tweezers; 100-mL beaker; 250-mL beaker

Safety: HCl is very corrosive—handle solutions carefully. If contact is made with skin, wash affected areas with plenty of water. Wear safety goggles at all times in the lab.

Waste Disposal: All reaction solutions and excess reagents can be flushed down the sink with plenty of water.

INTRODUCTION

As one of the three states of matter, gases have unique properties. They are readily compressible, and the volume occupied by a gas is affected dramatically by changes in temperature and pressure. Furthermore, the behavior of gases is governed by a well-defined set of laws which are applicable to all gases regardless of chemical identity.

Our understanding of the behavior of gases began with the work of Robert Boyle, who published his results in 1622. **Boyle's Law** states that the volume occupied by a gas (at constant temperature) is inversely related to the pressure of the gas. This relationship can be represented mathematically as

$$V \propto \frac{1}{P} \quad (1)$$

where V is the volume occupied by the gas at some pressure, P . Eq. (1) can also be written as

$$V = \frac{k_1}{P} \quad (2)$$

where k_1 is a proportionality constant.

More than a century later, the work of Jacques Charles and Joseph Gay-Lussac provided insight into the relationship between the temperature and volume of a gas. Gay-Lussac noted that the volume occupied by a gas decreases as temperature decreases, which led to the definition of **absolute zero** and the **Kelvin scale**. Temperatures on the Celsius scale can be converted to absolute (Kelvin) temperatures by Equation (3).

$$T = 273.15 + t \quad (3)$$

where T is the absolute temperature (in kelvins) and t is the temperature in Celsius. **Charles's Law** quantifies the relationship between temperature and volume, and states that the volume occupied by a gas (at constant pressure) is directly proportional to temperature. Charles's Law is represented mathematically as

$$V \propto T \quad (4)$$

or

$$V = k_2 T \quad (5)$$

where T is absolute temperature (in kelvins) and k_2 is a proportionality constant. Boyle's Law and Charles's Law can be rearranged as

$$\text{Boyle's Law: } PV = k_1 \quad (6)$$

$$\text{Charles' Law: } V/T = k_2 \quad (7)$$

We can combine Eqs. (6) and (7) into the **combined gas law**, which is given in Equation (8).

$$\frac{P \cdot V}{T} = k_3 \quad (8)$$

One advantage of the combined gas law is that it allows us to compare the temperature, pressure and volume of a gas sample under two different sets of conditions, where

$$\left(\frac{P_1 V_1}{T_1}\right) = \left(\frac{P_2 V_2}{T_2}\right) \quad (9)$$

If the volume of a gas sample is known at one set of conditions (T_1 , P_1), the volume can be calculated under another set of conditions (T_2 , P_2).

Intuitively, the volume occupied by a gas would be expected to increase in proportion to the number of moles of gas in the sample. This relationship, known as **Avogadro's Law** can be represented as

$$V = k_3 n \quad (10)$$

Equations (6), (7), and (10) can now be combined to yield the **ideal gas law**,

$$PV = nRT \quad (11)$$

where R is the proportionality constant known as the universal gas constant. The value of R is the same for all ideal gases. An **ideal gas** is defined as one whose behavior can be described exactly

by the gas laws presented here. Nearly all gases exhibit ideal behavior under standard conditions. Deviations from ideal behavior are observed, however, for polar gases at low temperatures and high pressures.

The value of R can be evaluated experimentally if all the other variables in Eq. (11) are known. In this exercise you will determine the value of R by measuring the volume of gas produced from a known reaction at a given temperature and pressure. The reaction used in this experiment produces $\text{H}_2(\text{g})$ by reaction of magnesium with hydrochloric acid:



By reacting a known mass of Mg in an excess of HCl, we can calculate the exact number of moles (n) of $\text{H}_2(\text{g})$ that will be produced. By running this reaction in a closed system (see Fig. 1) we can collect and measure the volume (V) of $\text{H}_2(\text{g})$ that is produced in a syringe. If atmospheric pressure (P) and temperature (T) are known, then Eq. (11) can be rearranged to solve for R :

$$\frac{PV}{nT} = R \quad (13)$$

The value (and units) of R will depend on the units used for P and V . The most common units of R are $\text{L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$.

Pre-Lab Questions

1. A student following the procedures in this lab exercise collected the following data:

mass Mg, g	0.0243
initial syringe volume, mL	0.8
final syringe volume, mL	26.3
barometric pressure, torr	748
temperature, K	293.5

Calculate the value of the universal gas constant, R .

2. What would be the volume of hydrogen gas produced by the reaction of 0.100 g of magnesium metal, collected at 750 torr and 298 K? Use the value of R found in question 1.
3. The results of this experiment are affected by the care with which each step is completed. Describe how each of the following errors would affect the calculated value of R .
- (a) The Mg slid into the HCl in the test tube before the system was sealed off.
 - (b) The room temperature increased by 5°C between the time that the final volume was measured and the temperature was recorded.
 - (c) The MgO coating was not removed before the Mg sample was placed in the HCl.
 - (d) The test tube was filled with 4 mL of 0.10 M HCl instead of 4 mL of 1 M HCl.

PROCEDURE

1. Obtain all the parts necessary to construct the gas collection apparatus shown in Figure 1. If not already assembled, assemble as instructed and have your lab instructor inspect it before proceeding. Be sure that the plunger assembly is lubricated and slides easily when pressure is applied.

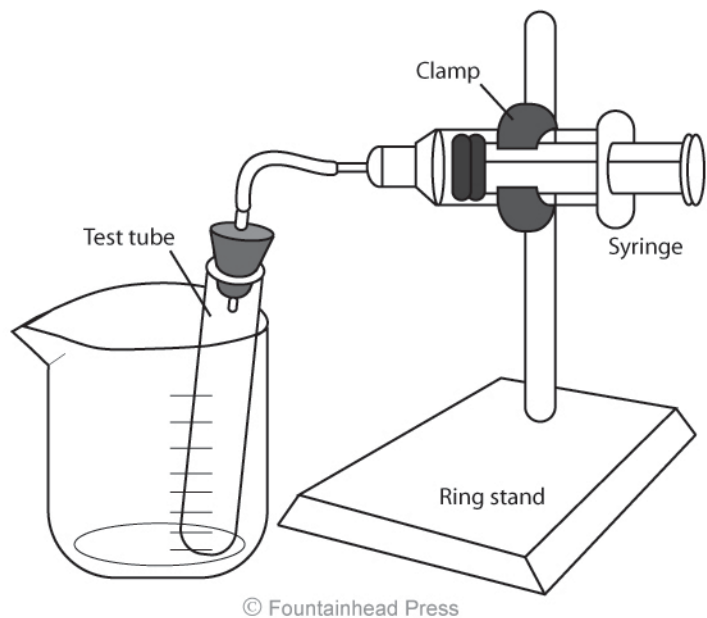


Figure 1. Apparatus for collection of $H_2(g)$.

2. Obtain 60 mL of 1 M HCl in a 100 mL beaker.
3. Prepare a 0.1 M HCl solution by placing 10 mL of the 1 M HCl obtained in Step 2 in a 250-mL beaker containing about 50 mL of DI water, and dilute to 100 mL. Label both beakers and retain these solutions for use later.
4. Obtain approximately 0.1 g of Mg ribbon. Magnesium is easily oxidized, so the Mg ribbon may have a white coating of MgO that must be removed before proceeding. Holding the ribbon carefully with a pair of tweezers, dip the Mg ribbon into the 0.1 M HCl solution. **Do not let the tweezers make contact with the acid solution.** When the Mg ribbon is clean, remove it from the HCl solution, rinse with distilled water and dry with an absorbent tissue.
5. Cut a portion of the clean Mg ribbon weighing approximately 0.02 g. While the exact mass is not critical, it is important that the mass be known precisely. **Record the mass of your Mg ribbon sample on your Data Sheet to four decimal places.**
6. Pipet 4 mL of the 1 M HCl solution into the test tube.

- Place your 0.02 g Mg sample in the test tube, with the tube tilted so that the Mg rests on the inside wall of the test tube but is not in contact with the HCl solution.
- Push the syringe plunger all the way in. Insert the rubber stopper in the test tube carefully so that the Mg ribbon does not slide into the HCl solution. If the plunger is lubricated appropriately to move freely, it should move out slightly when the rubber stopper is inserted in the test tube.
- Record the initial syringe volume reading to the nearest 0.1 mL on your Data Sheet.
- Tilt the test tube so that the Mg ribbon slides down into the 1 M HCl solution. As the reaction proceeds, H₂(g) will be produced and cause the plunger to move. Allow the reaction to proceed until all the Mg ribbon has reacted.
- Rotate the plunger without applying pressure to allow it to move freely. Let the reaction mixture equilibrate for 5 minutes, and rotate the plunger again.
- Record the final volume of the syringe on your Data Sheet to the nearest 0.1 mL.
- Record the barometric pressure and room temperature on your Data Sheet.
- Dispose of the reaction mixture in your test tube. Perform additional determinations as time permits by repeating Steps 4–13.

CALCULATIONS

Complete the following calculations for each determination, and record your results in the appropriate locations on the Data Sheet.

- Convert the mass of Mg to moles of Mg for each determination. Based on the stoichiometry of Eq. (12), this is also equal to the number of moles of H₂(g) produced.
- Calculate the volume of H₂(g) produced for each determination.
- Convert the observed Celsius temperature to Kelvin using Eq. (3).
- Since the H₂(g) was collected over water, the total pressure inside the test tube will be due to the partial pressure of the H₂(g) produced by the reaction as well as water vapor from the aqueous reaction solution. In the calculation of *R*, we must use only the partial pressure of H₂(g). We can correct for the contribution of water vapor using Eq. (14).

$$P_{H_2}(\text{atm}) = \frac{(\text{total pressure (torr)} - \text{vapor pressure } H_2O \text{ (torr)})}{760 \frac{\text{torr}}{\text{atm}}} \quad (14)$$

The vapor pressure of water as a function of temperature is given in Table 1.

5. Calculate R using the values of n , V , P , and T for each determination.
6. Calculate an average value of R using the results of all your determinations.

Table 1. Vapor pressure of water at temperatures from 18.0 to 25.0 °C.

Temp (°C)	v.p. H ₂ O (torr)		Temp (°C)	v.p. H ₂ O (torr)		Temp (°C)	v.p. H ₂ O (torr)
18.0	15.48		20.5	18.09		23.0	21.07
18.5	15.97		21.0	18.65		23.5	21.71
19.0	16.48		21.5	19.23		24.0	22.38
19.5	17.00		22.0	19.83		24.5	23.06
20.0	17.54		22.5	20.44		25.0	23.76

Data Sheet / Calculations

	<i>Determination</i>				
	1	2	3	4	5
mass Mg, g	_____	_____	_____	_____	_____
moles Mg, n	_____	_____	_____	_____	_____
syringe volume (initial), mL	_____	_____	_____	_____	_____
syringe volume (final), mL	_____	_____	_____	_____	_____
volume H ₂ (g), mL	_____	_____	_____	_____	_____
temperature (°C)	_____	_____	_____	_____	_____
absolute T (Kelvin)	_____	_____	_____	_____	_____
barometric pressure (torr)	_____	_____	_____	_____	_____
vapor pressure H ₂ O (torr)	_____	_____	_____	_____	_____
corrected pressure H ₂ (atm)	_____	_____	_____	_____	_____
Calculated R (L·atm/mol·K)	_____	_____	_____	_____	_____
Average R =	_____				

Sample Calculations:

Post-Lab Questions

- An evaluation of R was performed using the procedures outlined in this exercise. The barometric pressure was 735 torr and the temperature was 21.5°C . The volume of $\text{H}_2(\text{g})$ collected was 32.7 mL. The calculated value of R was $0.0817 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$.
 - How many grams of Mg metal were used in this determination?
 - If the correction for the vapor pressure of water had not been performed, what would be the calculated value of R ?
 - If the syringe volume had been incorrectly read resulting in a calculated H_2 volume of 30.6 mL, what would be the percent error in the calculated value of R ?
- Calculate the volume of one mole of hydrogen gas at 273 K and 760 torr. (Use a value of $R = 0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}$)
 - Would you expect the volume of one mole of oxygen gas under the same conditions to be the same, or different? Explain.
- Two moles of an ideal gas originally at 380 torr and 298 K are compressed at constant temperature to a final pressure of 680 torr. Use the ideal gas law to calculate the volumes of gas at the initial and final pressures.

Initial volume (@ 380 torr): _____

Final volume (@680 torr): _____