## Experiment 11

## A Determination of the Ideal Gas Constant

## Objectives

- To understand and practice gas laws.
- Experimentally to obtain a known amount of gas at known temperature, pressure, and volume and calculate the value of the ideal gas constant, $R$, using the $P V=n R T$.
- To understand a closed system.


## Introduction

The physical state of a sample of gas can be described completely with just four parameters: temperature $(T)$, pressure $(P)$, volume $(V)$, and number of moles $(n)$. Knowing any three of these variables allows us to calculate the fourth. The relationships between these parameters are known as the gas laws.

An ideal gas is a hypothetical sample of gas whose pressure-volume-temperature behavior is predicted accurately by the ideal gas equation. The ideal gas equation, $P V=n R T$, describes the relationship among the four variables $P, V, n$ and $T$. $R$ is called the ideal gas constant. Its value and units depend on the units in which $P$ and $V$ are expressed. Temperature must always be expressed in units of Kelvin in calculations using the ideal gas equation.

## Calculations

## Chemical Hazards

$3.5 \mathrm{~mol} / \mathrm{L}$ ( 3.5 M ) hydrochloric acid, HCl , is a strong acid. As a solution HCl is a corrosive substance and an irritant to skin and eyes.

- Skin contact can be detected by itching and/or redness. If you spill HCl on yourself, immediately flush the exposed skin area with running water in the sink or, flush the eyes in the eyewash station.
- Hydrogen gas is flammable and no open flames will be permitted during that part of the experiment.


## Chemical Waste

The remaining HCl solution should be very dilute at the end of the experiment because most of the acid has reacted. When finished, pour the remaining solution in the designated collection beaker.

## Procedure

1. Tare a glass vial on the balance. Weigh out between 0.6 and 0.8 g of magnesium turnings. A mass less than 0.6 g will likely give a poor result due to the small mass size while a mass of greater than 0.8 $g$ will probably result in your having to repeat the experiment (either one of the stoppers will "pop" off of a flask or you will exhaust your sample of water).

2. Get a 250 mL Erlenmeyer flask (Flask A) and put 40 mL 3.5 M HCl in it.
3. Get two 1000 mL Erlenmeyer flasks and a set of gas experiment tubing. Fill one flask (Flask B) up to the neck with tap water. You will need another student's help in order to properly set up the siphon between Flask B and Flask C as directed by your instructor. When you have the gas tubing properly set up, you will have tap water filled to the neck in Flask B, the stopper and glass tubing placed in Flask B, and the rubber tubing setting in Flask $C$. The pinch clamp (not shown) should still be on the tubing between Flask B and Flask C. Also, do not yet put the stopper in Flask A.
4. Carefully place the glass vial of magnesium in Flask $A$ using tweezers. Then place the stopper in Flask A without knocking over the vial.
5. To check the system for leaks, open the pinch clamp on the rubber tubing between Flask $B$ and Flask $C$ and then clamp it on the glass tubing coming out of the stopper in Flask B (in other words, it should not be clamped on the rubber tubing). A small amount of water will come out into Flask C but it should quickly stop. If the water does not stop, then you have a leak somewhere in your system. Check to ensure that the stoppers are tightly sealed in Flask A and Flask B. If it's still leaking, ask your instructor for help.
6. Before you do this step, you must make sure that the pinch clamp is not clamping the rubber tubing. Tightly hold the rubber stopper and Flask B in one hand while holding the rubber stopper and Flask A in the other hand. Gently rock Flask A only until the glass vial tips over. This will allow the magnesium to come in contact with the acid and begin the reaction.
7. The hydrogen gas produced will force water out of Flask B and into Flask C. Between 600 and 900 mL is usually pushed into Flask C at this point in the reaction. If not, ask your instructor for help because you may need to restart the reaction (usually only takes about 5 minutes to set it back up).
8. Once the reaction has slowed and no more water appears to be flowing from Flask B to Flask C, put a thermometer in Flask C. You will now need to wait approximately 45 minutes for the gas and water to equilibrate to room temperature.
9. Put the pinch clamp back on the rubber tubing between Flask B and Flask C. Once you've done this no more water should flow between those two flasks. Take the temperature of the water in Flask C. Since the entire system including the hydrogen gas and the water should be about room temperature, the gas temperature is approximately the same as the water temperature.
10. Disassemble the gas tubing from the flasks and use a 1000 mL graduated cylinder to measure the water collected in Flask C.
11. Dispose of the excess HCl solution in the collection beaker. Be careful that you don't break the small glass vial that previously held the sample of magnesium. Set the vial in the "Wet Vial" container.
12. Clean your lab station and complete the calculations and analysis of the class's results.

The relationship between the vapor pressure of water and temperature is modeled by a rather complicated mathematical formula that can be accessed at the following web link (use the "Calculate the Vapor Pressure of Water" data entry section)-
http://chemistry.alanearhart.org/Lab/CHEM1090/exp11.html
$\qquad$ Date $\qquad$
$\qquad$

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1. Data:

1a. Mass of Magnesium $\qquad$
1b. Volume of Water in Flask C

1c. Water Temperature
1d. Barometric Pressure

1e. Vapor Pressure of Water
$\qquad$
$\qquad$
$\qquad$
2. Calculations (the objective is to calculate the value of $R$ therefore you will not use the value of $R$ in this section):

2a. Write the chemical equation that produced hydrogen gas from solid magnesium and hydrochloric acid:

2b. Calculate the volume of hydrogen gas produced in liters:

2c. Calculate the pressure of hydrogen gas (in Torr) using $P_{\text {total }}=P_{\text {hydrogen gas }}+P_{\text {water vapor: }}$ :

2d. Calculate the pressure of hydrogen gas in atmospheres, atm:

2e. Calculate the absolute (Kelvin) temperature of the hydrogen gas:

2f. Calculate the moles of hydrogen gas produced using the mass of magnesium:

2 g . Calculate the value of the gas constant R in units of $\frac{\mathrm{Latm}}{\mathrm{mol} \mathrm{K}}$ :

2 h . Determine the value of the ideal gas constant from your data using the lab resources webpage. The value of the ideal gas constant should differ no more than $\pm 2$ in the last significant figure from your value in 2 g . Write your value below:

## Analysis

1. Which percent error do you expect to be lower: Your value or the class's value? Explain why.
2. Write the class's values for the ideal gas constant.
3. Determine the average and the standard deviation for the class's values and write them to four decimal places (include units).

Average $\qquad$
$\qquad$
4. Calculate the percent error using your calculated value of the ideal gas constant as the EV and $0.082057 \frac{\mathrm{~L} \mathrm{~atm}}{\mathrm{~mol} \mathrm{~K}}$ as the SV.
5. Calculate the percent error using the class's average value of the ideal gas constant as the $E V$ and the same value for the $S V$ as the previous question.
6. Which percent error was actually lower: Your value or the class's value? Does this match your expectation? Why or why not?
7. How do you know that magnesium was the limiting reagent for this reaction?
8. We have been using the rules for keeping track of significant figures as set up during the first lab and the first week of lecture. Statistics can be also used to determine how to write your result.

8a. Take the standard deviation determined earlier and round it to one significant figure:
Standard Deviation $\qquad$

8b. Take the class's average value for the ideal gas constant and round to the same number of decimal places as the standard deviation:

Average $\qquad$
8c. How many significant figures are in your average?

