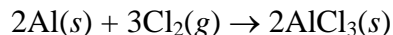


## Experiment 3

### Limiting Reactants

**Introduction:** Most chemical reactions require two or more reactants. Typically, one of the reactants is used up before the other, at which time the reaction stops. The chemical that is used up is called the *limiting reactant* while the other reactant is present in *excess*. If both reactants are present in exactly the right amount to react completely, without either in excess, the amounts of reactants are said to be in a *stoichiometric ratio* to each other. The stoichiometric ratio is the mole ratio of the reactants, or reactants to products, as determined by the coefficients in the balanced chemical equation. Since the limiting reactant will determine the amount of product that can be produced during a reaction, it is important to be able to calculate which reactant is the limiting reactant. There are several ways to do this, but each starts with a balanced chemical equation so that the stoichiometry of the reaction is known.

For example, the balanced equation for the synthesis of aluminum chloride shows that two moles of aluminum react with three moles of chlorine gas to produce two moles of the aluminum chloride product:



If the reaction is carried out with 50.00 g of aluminum and 150.00 g of chlorine, the reaction will be observed to stop when the chlorine runs out, *even though there is a greater mass of chlorine*. Thus, chlorine is the limiting reactant. To see why this is so, the mass of each reactant must first be converted to moles using the formula masses. This is done by the first conversion factor in the example below. The second conversion factor multiplies the moles of each reactant by the stoichiometric ratio relating moles of product to moles of reactant. Finally, the third conversion factor changes moles of product into grams of product formed from the given amounts of each reactant.

This method of determining limiting reactants allows each reactant to be related to the same product to determine which will result in the **least amount of the product formed**. The reactant that does this is the limiting reactant.

$$50.00 \text{ g Al} \times \left( \frac{1 \text{ mol Al}}{26.9815 \text{ g Al}} \right) \times \left( \frac{2 \text{ mol AlCl}_3}{2 \text{ mol Al}} \right) \times \left( \frac{133.34 \text{ g AlCl}_3}{1 \text{ mol AlCl}_3} \right) = 247.1 \text{ g AlCl}_3$$

$$150.00 \text{ g Cl}_2 \times \left( \frac{1 \text{ mol Cl}_2}{70.906 \text{ g Cl}_2} \right) \times \left( \frac{2 \text{ mol AlCl}_3}{3 \text{ mol Cl}_2} \right) \times \left( \frac{133.34 \text{ g AlCl}_3}{1 \text{ mol AlCl}_3} \right) = 188.1 \text{ g AlCl}_3$$

As illustrated above, 150.00 g of  $\text{Cl}_2$  produces only 188.1 g of  $\text{AlCl}_3$ , as compared to 247.1 g of  $\text{AlCl}_3$  from 50.00 g of Al. Therefore,  $\text{Cl}_2$  is the limiting reactant. *Note that this result would also be obtained by comparison of the moles of  $\text{AlCl}_3$  produced.*

It is particularly beneficial to determine the limiting reactant this way if the *theoretical yield* is also desired. The theoretical yield is the amount of product that will be obtained if all the limiting reactant is converted to product. The theoretical yield must therefore be calculated based upon the limiting reactant, as no additional product can be formed once it has been used up. The limiting reactant is related to the product using the stoichiometry of the balanced equation. In the example above, since  $\text{Cl}_2$  is the limiting reactant and it could form 188.1 g of  $\text{AlCl}_3$  product, that will be the theoretical yield for the reaction.

Of course, in practice this yield is rarely attained. There are many potential losses of the product during its collection after a reaction and non-product forming side reactions usually occur as well. The actual experimentally measured yield of the product is expressed as a percentage of the theoretical yield and is called the *actual percent yield* or just *percent yield*.

$$\% \text{ Yield} = \left( \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) \times 100$$

If, for the example above, reacting 50.00 g of Al with 150.00 g of  $\text{Cl}_2$  resulted in an actual yield of  $\text{AlCl}_3$  of just 135.5 grams, the percent yield would be 72.04%.

$$\% \text{ Yield} = \left( \frac{135.5 \text{ g}}{188.1 \text{ g}} \right) \times 100 = 72.04\%$$

Both the theoretical yield and the actual yield must be in the same units so that the % yield is a unitless quantity. However, these yield units need not be only grams; the amount can also be expressed in moles, volume, or concentration of product (See the following section).

**Molarity:** The amount of solute present in a given volume of solution is called *concentration*. There are several units for concentration but the most often used is *molarity* in which the amount of solute is given in moles and the volume in liters of solution. A one molar solution is defined as a solution having a concentration of 1 mole of solute per liter of solution and is abbreviated 1 *M*.

$$\text{Molarity (M)} = \frac{\text{mol solute}}{\text{L solution}}$$

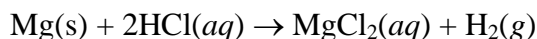
When solutions have a concentration given in molarity, the moles of solute in any sample of the solution is directly related to the volume of that solution. For example, if 20.2 mL of a 1.045 *M* reactant solution is used for a reaction, the moles of that reactant can be calculated from the definition of molarity.

$$1.045 \text{ M} = \left( \frac{1.045 \text{ mol}}{\text{L}} \right) = \left( \frac{1.045 \text{ mol}}{1000 \text{ mL}} \right) \quad \text{so that} \quad \left( \frac{20.2 \text{ mL}}{\text{L}} \right) \times \left( \frac{1.045 \text{ mol}}{1000 \text{ mL}} \right) = 0.0211 \text{ mol}$$

Therefore 20.2 mL of the solution contains 0.0211 mol of the solute.

### Reaction to Produce the Gas, H<sub>2</sub>:

The objective of this experiment is to discover how varying the relative amounts of reactants affects the amount of product produced in a chemical reaction, and thus confirm the concept of limiting reactants. You will use varying **masses of magnesium** metal while holding the hydrochloric acid concentration and volume constant and observing the **pressure of hydrogen gas** produced during each run. Under the conditions of the experiment, the pressure of H<sub>2</sub>(g) is directly related to the amount in moles of H<sub>2</sub>(g) produced by the reaction. The balanced equation for the reaction is:



To save time and achieve more results for use in the analysis, this experiment will be a class data-pooling project. You will run trials on two lengths of magnesium ribbon (each done in duplicate), and then you will combine your data and results with fellow classmates who will run trials on different amounts of magnesium. Your instructor will construct the table shown below on the chalkboard. Two lengths of Mg ribbon will be assigned to each team starting with team 1 and working down the list. After teams have determined the **average of the mass of the Mg samples and the average of the pressure of H<sub>2</sub> produced for each of their two lengths**, they will enter the values in the table. Since the volume of HCl is the same for all runs at 3.0 mL, it is not necessary to enter that value.

Team #	length (cm)	mass (mg)	P(H <sub>2</sub> ) (torr)	length (cm)	mass (mg)	P(H <sub>2</sub> ) (torr)
1	1.0			4.0		
2	2.0			5.0		
3	3.0			6.0		
4	1.3			4.3		
5	2.3			5.3		
6	3.3			6.3		
7	1.7			4.7		
8	2.7			5.7		
9	3.7			6.7		
10	1.5			4.5		
11	2.5			5.5		

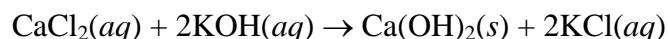
Students must transfer the data entered by all teams in the above table into a similar table in their lab notebooks before getting their instructor's initials in their lab notebooks.

You will need to know that when collecting a gas over an aqueous solution, in this case H<sub>2</sub> over aqueous HCl, the collected gas contains H<sub>2</sub>O vapor. The gas in the flask at the end of the reaction will also contain the air initially in the Erlenmeyer flask. The total gas pressure measured by the Vernier probe will thus be the sum of the three component gases in the flask. Therefore the pressure of the hydrogen gas must be calculated using the asterisked equation shown under the data table. The greater the pressure of H<sub>2</sub> gas calculated for reactions carried out in the same way, the greater the amount of H<sub>2</sub> produced.

You should keep this question in mind while performing the experiment: If, in a series of experiments, you use increasing amounts of magnesium metal while holding the amount of HCl constant, will the limiting reactant for the reaction change and, if so, how will this change affect the pressure of H<sub>2</sub> gas produced in the series of experiments?

### Reaction to Produce the Solid, Ca(OH)<sub>2</sub>:

In a separate reaction of calcium chloride and potassium hydroxide (performed in Part B of this experiment), you will calculate the limiting reagent, theoretical yield, and percent yield after collection of the calcium hydroxide product. The balanced reaction is shown below.



After separating the Ca(OH)<sub>2</sub>(s) formed from the reaction mixture, you will be able to determine the mass of Ca(OH)<sub>2</sub>(s) and compare it to the theoretical yield to find the percent yield for the reaction.

**Pre-Laboratory Notebook:** Provide a title, purpose, both balanced chemical reactions, a brief summary of the procedure, and table of reagents and products such as the following:

#### Table of Chemicals Used:

Compound	Formula Mass, (g/mol)	Hazards
Magnesium, Mg		
Hydrochloric acid, HCl		
Magnesium chloride, MgCl <sub>2</sub>		
Hydrogen, H <sub>2</sub>		
Calcium chloride, CaCl <sub>2</sub>		
Potassium hydroxide, KOH		
Calcium hydroxide, Ca(OH) <sub>2</sub>		
Potassium chloride, KCl		

Also, before coming to lab, create a data table in your notebook similar to the one below:

Run	1	2	3	4
Amount of Mg (mg)				
Amount of HCl (mL)				
Highest Pressure, P <sub>total</sub> (torr)				
Initial Pressure, P <sub>init</sub> (torr)				
Temperature (°C)				
P <sub>H<sub>2</sub>O</sub> (torr) from table below				
P <sub>H<sub>2</sub></sub> (torr)*				

$$* P_{\text{H}_2} = P_{\text{total}} - P_{\text{init}} - P_{\text{H}_2\text{O}}$$

## Vapor Pressure of Water at Various Temperatures

Temperature (°C)	Water vapor pressure (torr)
20.0	17.5
21.0	18.7
22.0	19.8
23.0	21.1
24.0	22.4
25.0	23.8
26.0	25.2

### Equipment:

Vernier LabPro	125 mL Erlenmeyer flask
Vernier gas pressure sensor	Magnesium ribbon
TI -84 calculator	1.0 M HCl
2, 100 mL beakers	100 mL beaker (for HCl solution)
Steel wool	Liquid funnel and filter paper
Stir rod	Ring stand and ring

### In Lab Experimental Procedure:

Note: Work in pairs.

#### Part A: How Much Is Too Much?<sup>1</sup>

- Set up the LabPro system in DATAMATE according to the first two sections of instructions in the Vernier tutorial. Connect the gas pressure sensor into channel 1 of the LabPro. The main screen of DATAMATE should show a pressure in the lab of ~600-620 torr. Your instructor will write the lab temperature and the actual lab pressure as measured by a barometer on the chalkboard. The unit shown on the calculator is MMHG (mm of mercury), which is identical to torr. If the pressure reading is not ~600-620 torr (MMHG), check with your Instructor.
- If the pressure units are other than MMHG, select 1:SETUP from the main screen. Press  to select CH 1. Then select 2: Gas Pressure (MMHG). Select 1: OK twice to return to the main screen.
- Using steel wool, polish a length of Mg ribbon which is about equal to the sum of the four lengths assigned to you. From this polished strip, cut your four assigned samples of magnesium ribbon (two each of the two assigned lengths). Then cut each sample into ~1 cm long strips and weigh them in tared weigh boats. Record the exact mass (in mg units) of each sample in the data table you created in your lab notebook. (1 cm is about equal to 9 mg, so a 5cm sample should weigh ~45 mg). You can keep the Mg strips in the weigh boats until ready to proceed with the reactions, but keep track of which weigh boat contains which sample!
- Using a clean 125 mL Erlenmeyer flask, set up the apparatus as shown in Figure 1 below with the first sample of Mg in the flask, but do not add HCl into the syringe at this point. Ensure that all connections are tightly secured, but **do not over tighten screw on connectors**. The connection from the stopcock to the stopper feed through

must not be disconnected. Downward pressure should be applied to ensure that the stopcock is securely attached to the feed through and that the feed throughs are securely in place in the stopper. When the stopper is firmly in place, it should not be more than halfway below the lip of the flask

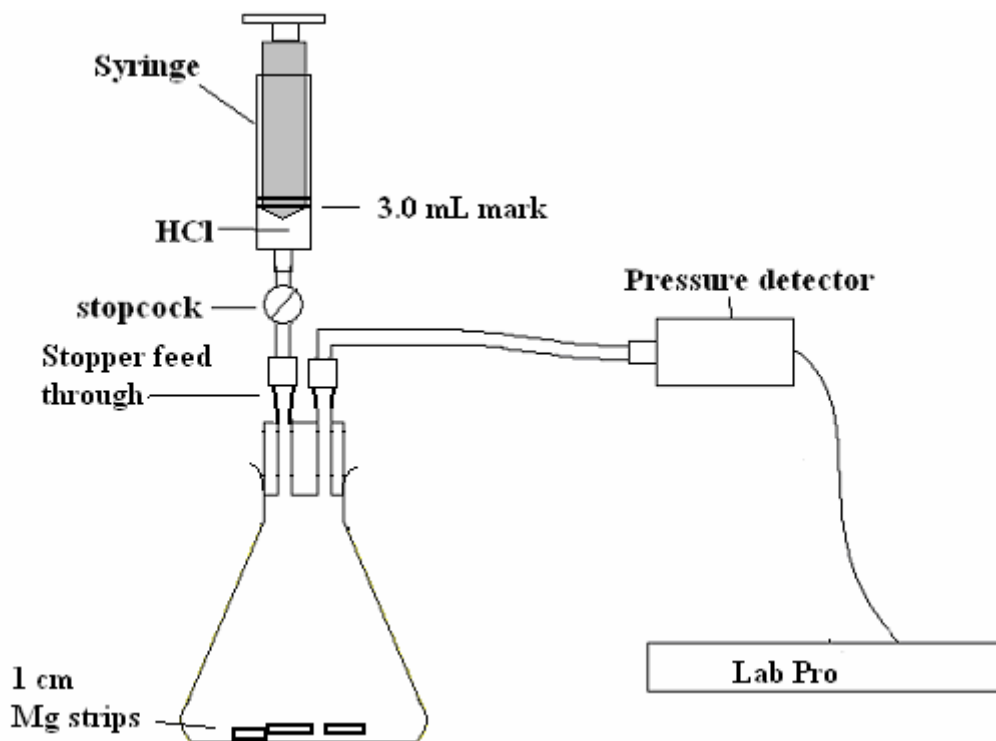


Figure 1. Apparatus for reacting Mg and HCl to form H<sub>2</sub>(g).

5. With the completely depressed syringe screwed onto the stopcock, retract the plunger to the 20 mL mark to remove air from the flask and then close the stopcock. The pressure should be reduced to about 520 torr.
6. Unscrew the stopcock and depress the plunger completely. Screw the stopcock back on to the syringe. Then **OPEN the stopcock** and withdraw another 20 ml air from the flask, and **then CLOSE the stopcock**. This should lower the pressure to ~450 torr. Repeat this process until the pressure in the flask is about 350 torr. With the stopcock remaining closed, check that the pressure is not increasing more than about 3 torr in 3 minutes, indicating reasonably secure seals.
7. Unscrew the syringe and draw up 3.0 mL of HCl solution from the beaker into the 20 mL syringe. Note: It is good practice to add more than 3 mL initially. If there are significant air bubbles in the syringe, point the syringe upwards to get the air at the tip of the syringe and carefully discharge a small amount of the air and acid into a paper towel. Then point the syringe downward and depress the plunger, releasing excess HCl back into its beaker, until the volume in the syringe is 3.0 mL. Read the volume at the front edge of the inside black ring on the piston of the syringe as shown in Figure 1. Connect the syringe onto the two-way valve on the white stopper with a twisting motion to tighten it in place.

8. Set up the data collection mode. Each run should last just over eight minutes, with a data sample taken every five seconds. Select SETUP from the main screen. To select MODE, press  $\blacktriangle$  or  $\blacktriangledown$  until the cursor is to the left of MODE and press  $\text{ENTER}$ . Select TIME GRAPH from the SELECT MODE menu. Select CHANGE TIME SETTINGS. Type “5”, for the time between samples in seconds, and press  $\text{ENTER}$ . Type “100”, for the number of samples, and press  $\text{ENTER}$ . Select OK twice to return to the main screen.
9. **Before adding the HCl to the flask**, select START from the main screen on the calculator to begin data collection. After 20 seconds, open the stopcock, press the plunger smoothly to add all of the 3.0 mL of HCl solution to the flask, and close the stopcock immediately.
10. Check that the stopper remains in place while swirling the apparatus to ensure mixing of the reactants. Wait for the beep from the Vernier system to signify the end of data collection. In a few seconds following the end of data collection, the screen will display the list of channels. With the cursor arrow at CH 1 – Pressure, press  $\text{ENTER}$  and a plot of pressure vs. time will appear. Use the cursor arrow buttons to scan through the plot and record the pressure (y value) at the start of the plot (before addition of HCl); this is  $P_{\text{init}}$ . Then locate and record the maximum pressure following reaction; this is  $P_{\text{total}}$ . Enter these values into your data table in your lab notebook.
11. Unscrew the syringe from the stopper and open the stopcock to return the flask to atmospheric pressure. Press  $\text{ENTER}$  to clear the Lab Pro and press 1-Main screen. Then repeat the process for the next sample of Mg strips starting at step 4.
12. **In your lab notebook, directly under your data table**, show the calculations for the pressures of  $\text{H}_2$  for each run. Average the masses of the two equal strips of magnesium used and the two calculated pressures of the  $\text{H}_2$ . Record these averages (to the nearest mg of Mg and torr of  $\text{H}_2$ ) below your data table and on the chalkboard for both assigned masses of magnesium. The pool of class data should be recorded in a table in your notebook that resembles the table on the chalkboard.
13. Once you have obtained data from the entire class, plot the pressure of the hydrogen gas (y-axis) vs. the mass of Mg (x-axis) on Microsoft Excel. You may discard points on the graph that differ significantly from the trend observed for the plot. Add appropriate titles and units to your Excel plot, and then print a copy for submission with your lab report. This graph will be used for data analysis in the results and discussion session.

### Part B: Percent Yield of $\text{Ca}(\text{OH})_2$

1. Measure out 0.490–0.510 g of each reactant,  $\text{CaCl}_2 \cdot 2\text{H}_2\text{O}$  and KOH, in weigh boats. Note that while  $\text{CaCl}_2$  is the reactant in the equation for formation of  $\text{Ca}(\text{OH})_2$ , it is the hydrate,  $\text{CaCl}_2 \cdot 2\text{H}_2\text{O}$ , which is the source of the  $\text{CaCl}_2$ . It has a formula mass of 147.02 g/mol, and clearly there is one mole of  $\text{CaCl}_2$  for each mole of  $\text{CaCl}_2 \cdot 2\text{H}_2\text{O}$  used. Therefore when finding the number of moles of  $\text{CaCl}_2$  for the purpose of determining the limiting reactant, **you must use the mass and formula mass of  $\text{CaCl}_2 \cdot 2\text{H}_2\text{O}$ .**
2. Record the exact mass of the reactants used in your laboratory notebook. Transfer all of the solid in each weigh boat to separate 100 mL beakers, using de-ionized water. Add

additional de-ionized water to the 30 mL mark on the beaker. Swirl each beaker until the solid reactants have dissolved completely.

3. Pour the two aqueous solutions together and mix well with a stir rod. A white precipitate will form almost immediately, but allow the solution to sit for a few minutes to ensure a complete reaction.
4. Measure and record the mass of a piece of filter paper. Collect the  $\text{Ca}(\text{OH})_2$  precipitate onto the paper by performing a gravity filtration.
5. Remove the filter paper from the funnel and place it in a labeled weigh boat to dry. After at least one day, you can weigh the dried filter paper plus product during any ongoing laboratory period prior to your next scheduled lab period.

### Results and Discussion:

#### Part A: How Much Is Too Much?<sup>1</sup>

1. Is the limiting reagent the same throughout the sequence of experiments with different masses of Mg? Explain your answer based on the Excel plot of  $P(\text{H}_2)$  vs. mass Mg.
2. Initially, what happens to the pressure of hydrogen gas collected as the mass of magnesium metal is increased? Would you expect this trend continue indefinitely? Explain your answer based on the Excel plot of  $P(\text{H}_2)$  vs. mass Mg.
3. Which starting reactant is limiting when the mass of magnesium metal is 50 mg? Which starting reactant is limiting when the mass of magnesium metal is 20 mg? Explain your answer based on the Excel plot of  $P(\text{H}_2)$  vs. mass Mg.
4. When using 3.0 mL of 1.0 M HCl to react with magnesium metal, at what mass of magnesium (in mg) are the two reactants in a stoichiometric ratio to each other? Show your calculation to determine this mass of Mg. What happens to the class plot at this particular mass of magnesium?
5. Would it be possible to run a chemical reaction with no single limiting reagent? Why or why not?



## Lab Report Outline for Limiting Reactants

Graded Sections	Percent of grade
Pre Lab	10
In Lab	
Data table for runs 1-4	10
Calculations of P(H <sub>2</sub> ), torr	5
Calculations of averages of masses and P(H <sub>2</sub> ), torr	5
Post Lab	
A. Mg + HCl reaction:	
Results and Discussion questions 1-5	25
Microsoft Excel graph	10
B. CaCl <sub>2</sub> + KOH reaction: ( <b>First restate your key data</b> )	
DATA: Mass KOH: _____g, Mass CaCl <sub>2</sub> ·2H <sub>2</sub> O: _____g,	
Mass dry filter: _____g; Mass dry filter plus Ca(OH) <sub>2</sub> : _____g;	5
1. Calculation of moles of CaCl <sub>2</sub> (recall CaCl <sub>2</sub> ·2H <sub>2</sub> O was used) and KOH <input type="checkbox"/> <input type="checkbox"/>	10
2. Determination of limiting reagent and grams Ca(OH) <sub>2</sub> in theoretical yield <input type="checkbox"/>	10
3. Calculation of actual yield of Ca(OH) <sub>2</sub> in grams. <input type="checkbox"/>	5
4. Calculation of actual % yield. <input type="checkbox"/>	5

1. "Experiment #3 How Much is Too Much?" Ram Lamba, Shiva Sharma, and Baird W. Lloyd, in "Constructing Chemical Concepts through a Study of Metals and Metal Ions," *J. Chem. Ed.* **74** (1997), 1095-1099.

