A SourceBook Module

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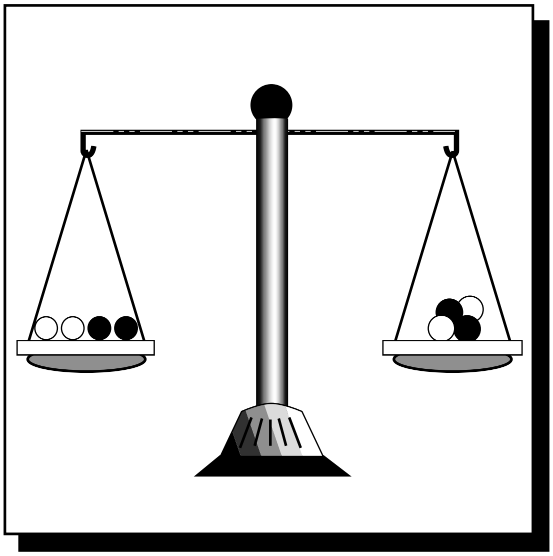
STOICHIOMETRY



ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

The big question in chemistry laboratory work is often “how much?” This same question pervades the realm of chemical research and industrial production. Since the work of early scientists such as Lavoisier, one of the first to make careful, quantitative observations about reactions, this question has centered around the understanding of the Law of Conservation of Mass.

With the development of the mole concept, the basis for quantitative evaluation of chemical reactions has been associated with the balanced chemical equation, which derives its validity from the Law of Conservation of Mass. The study of the quantitative relationships in chemical reactions is commonly known as **stoichiometry**, deriving its name from the Greek words *stoicheion* for element and *metrein* for measure.

Since the work of Dalton, quantitative relationships of atoms combining with each other to form compounds of definite ratios has continued to organize chemists’ approach to the study of stoichiometry. Chemists today use equation coeffiecients as the basis for quantitative study of chemical change. The equation coefficients represent moles of substances reacting and moles of substances produced (assuming the reaction goes to completion). This is very convenient since the amount of a substance can be expressed in terms of moles.

Frequently chemical reactions involve solutions. Using molarity to express the concentration of a solution facilitates stoichiometric calculations. For gases, the ideal gas law can be used to relate volume to moles. (This law is acceptable for most laboratory-scale reactions but may not be suitable for industrial-scale processes taking place at high pressure.) The important principle is that equation coefficients are the basis for stoichiometric calculations, and the means by which the number of moles involved in the reaction is determined is secondary.

When the concepts of limiting reactant and percentage yield are included, the study of stoichiometry begins to have added significance in industrial and research applications. It is important, however, not to overstate the idea that the limiting reactant “controls” the extent of a reaction. Most reactions encountered in beginning chemistry courses have equilibrium positions displaced toward products. However, all reactions do not “go to completion.” In such reactions, the limiting reactant is not the controlling factor; the percent yield may be much less than 100%. (Factors affecting the position of equilibrium are discussed in the *Chemical Equilibrium* module; kinetic factors are considered in the *Rates of Reaction* module.)

PLACE IN THE CURRICULUM

In general and honors courses stoichiometry is discussed early in the course, usually after the mole concept discussion as an application example. For students with weak mathematics skills, stoichiometry may be delayed; however, balancing equations, discussed in this module, may be presented earlier.

CENTRAL CONCEPTS

1. The coefficients of a chemical equation are the basis for interpreting and predicting quantitative relationships in chemical reactions (stoichiometry).
2. Stoichiometric relationships are based on the mole concept.
3. The reactant present in the smallest stoichiometric amount (the limiting reactant) determines the amounts of other reactants consumed and the amount of product formed.
4. The actual yield of product in a reaction may be less than the theoretical yield.

RELATED CONCEPTS

1. Mass, volume
2. Formulas (empirical and molecular)
3. Atoms, molecules, formula units
4. Accuracy and precision
5. Gas behavior (particulate nature of matter, relationship of volume to moles of gas)
6. The mole
7. Chemical equations
8. Reactant, product
9. Percent
10. Conservation of matter
11. Solution concentration (molarity)

RELATED SKILLS

1. Exponential numbers
2. Significant figure manipulation
3. Problem solving strategies (dimensional analysis, *etc.*)
4. Algebraic manipulations
5. Graphing
6. Observation
7. Mass, volume and temperature measurements
8. Use of burner/hot plate

PERFORMANCE OBJECTIVES

After completing a study of stoichiometry, students should be able to:

1. explain how the total mass before reaction is related to the total mass after reaction (whether or not the reaction goes to completion).
2. explain how a balanced chemical equation relates to the law of conservation of mass.
3. balance chemical equations and explain how the coefficients are related to the amount (in moles) of reactants and products.
4. determine the amount (in moles) of any substance in a reaction, given the amount (in moles) of any other substance.
5. determine the mass of any substance in a reaction, given the amount (in moles) of any other substance.
6. determine the mass of any substance in a reaction, given the mass of any other substance.
7. explain what is meant by "limiting reactant."
8. identify the limiting reactant in a reaction and calculate the theoretical yield of product(s) and the quantity of other reactant(s) consumed or unreacted.
9. explain the meaning of percent yield and explain why actual yield is often less than theoretical yield.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Stoichiometry of a Reaction and Limiting Reactant

Introduction

A chemical reaction can be described by an equation. The equation shows the relationship of the moles of reactants to moles of products. Since matter is neither created nor destroyed in chemical reactions (Law of Conservation of Mass), the equation must show equal numbers of atoms of each element in both reactants and products. Such an equation is said to be "balanced."

Stoichiometry is the study of quantity relationships in chemical reactions. The stoichiometry of a reaction is based on the mole. The coefficients in the balanced equation represent the relative numbers of moles of substances that react or are formed in the reaction.

For practical reasons, the quantities of the reactants used in a reaction are generally not in the same mole proportions as the balanced chemical equation. It is important to be able to predict how much product will form under these conditions.

Purpose

To demonstrate that substances react in proportions related to the number of moles of reactants. To show that if the reactants are not present in stoichiometric proportions, the substance present in smallest stoichiometric amount (limiting reactant) will determine the amount of product(s) formed.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Carefully clean up any spills.
3. Exercise caution in lighting the burner. Loose hair should be tied back. Hair ignites easily especially if hair spray is used.
4. Always exercise caution when handling glassware especially if it is hot. Hot and cold glass look the same!

Procedure

1. Weigh accurately a clean, dry 50- or 100-mL beaker. Record the mass. Add to the beaker the assigned quantity (0.2-0.5 g) of iron filings. Reweigh accurately the beaker and contents. Record the mass.
2. Weigh accurately a 150- or 250-mL beaker. Add approximately 3.0 g $CuSO_4 \cdot 5H_2O$. Reweigh accurately and record the mass of beaker and $CuSO_4 \cdot 5H_2O$. Add approximately 50 mL water (preferably distilled water) to the beaker.

   Warm the beaker and stir carefully to dissolve the solid. (Do not boil the solution.) Record the solution's color.

   Alternate procedure: Using a graduated cylinder, measure carefully approximately 25 mL 0.500 M $CuSO_4$. Record accurately the volume. Add the solution to the weighed 150- or 250-mL beaker.

3. Remove the burner under the beaker (or remove the beaker from the hot plate) and carefully pour the iron filings into the $CuSO_4$ solution a little at a time. After each addition of iron filings, wait for the reaction to subside. Record your observations.
4. To speed up the reaction, *gently* heat the beaker. The reaction should be complete when the solution reaches the boiling point. *Do not boil the liquid*, you may lose some of your product. Record your observations.
5. When the reaction is complete (no iron filings remain; look from below if you have a problem seeing whether iron filings remain), allow the solid to settle. Let the beaker cool until it is safe to touch. Add 1-2 drops of a detergent solution to prevent the solid particles from floating on the surface of the liquid. Carefully decant into a beaker (pour off the liquid leaving the solid behind). Record the appearance of the decantate (the liquid poured off). You will not be able to pour off all the liquid, but pour off as much as possible without losing any of the solid. (If you accidentally pour some of the solid off with the decantate, you can carefully pour the decantate into another beaker, wash the particles back into the beaker with the solid reaction product and try decanting again.)*Dispose of the decantate as your teacher directs.*
6. Wash the solid by adding 10 mL water (preferably distilled water). Swirl *very* gently. (Be *very* careful not to spill the solid and ruin your results!) Let the solid settle. Decant (as described in Step 5). Continue the washing procedure several times until the wash water is colorless. After the last washing decant as much water as possible so drying will not take so long.
7. Dry the solid in a drying oven. Alternatively, place the beaker where it can remain undisturbed; allow to dry overnight. (The beaker may also be placed under a heat lamp to evaporate the water.) Do not leave the solid in a drying oven overnight. If the solid is allowed to stay at an elevated temperature for a long period of time, it may react with oxygen in the air and alter the results.
8. Cool the beaker. Weigh the beaker and contents accurately and record the mass.
9. Thoroughly wash your hands before leaving the laboratory.
10. Prepare a data table like the one shown.

Data Table

1. Mass of 50-mL beaker ____________________
2. Mass of 50-mL beaker and iron filings ____________________
3. Mass of iron filings ____________________
4. Mass of 150-mL beaker ____________________
5. Mass of 150-mL beaker and $CuSO_4 \cdot 5H_2O$ ____________________
6. Mass of $CuSO_4 \cdot 5H_2O$ ____________________
7. Mass of 150-mL beaker and Cu ____________________
8. Mass of Cu ____________________

Data Table for Alternate Procedure:

1. Mass of 50-mL beaker ______________________
2. Mass of 50-mL beaker and iron filings ______________________
3. Mass of iron filings ______________________
4. Volume 0.500 M $CuSO_4$ ______________________
5. Mass of $CuSO_4$ ______________________
6. Mass of 150-mL beaker ______________________
7. Mass of 150-mL beaker and Cu ______________________
8. Mass of Cu ______________________

Part I. Data Analysis: Stoichiometry of the Reaction

1. Calculate moles of Fe used from the mass of Fe used and the molar mass of Fe.
2. Calculate moles $CuSO_4$ used. Calculate moles of $CuSO_4 \cdot 5H_2O$ from the mass of $CuSO_4 \cdot 5H_2O$ used and the molar mass of $CuSO_4 \cdot 5H_2O$. Since one mole of $CuSO_4 \cdot 5H_2O$ contains one mole of $CuSO_4$, moles of $CuSO_4 \cdot 5H_2O$ equals moles of $CuSO_4$.

   [Alternate procedure: Moles of $CuSO_4$ is calculated from the molarity and volume of $CuSO_4$: Moles = Molarity x Volume (in liters)]
3. Calculate moles Cu formed in the reaction from the mass of Cu obtained and the molar mass of Cu.
4. Tabulate the mole data for the class. Plot a graph of moles Cu *vs.* moles Fe. (Graph moles Cu on the y-axis; moles Fe on the x-axis.) Draw the best straight line through the data points.
5. Determine the slope of the line. The slope of the line gives the mole ratio (mol Cu/mol Fe). Remembering that the coefficients in the balanced equation should be integers, use this relationship between mol Cu and mol Fe to write a balanced equation for the reaction of $CuSO_4$ and Fe to form Cu and $FeSO_4$.

Part I. Implications and Applications

1. How would your experimental value for moles Cu be affected in the following situations?
   a. If 4 g $CuSO_4 \cdot 5H_2O$ were added instead of 3 g $CuSO_4 \cdot 5H_2O$?
   (If 30-mL 0.50 M $CuSO_4$ were added instead of 25-mL 0.50 M $CuSO_4$?)
   b. If the solid did not settle before decanting and some solid product was poured off during decanting (in Steps 5 and 6)?
   c. If the solid product was not dried thoroughly and weighed wet?
   d. If the beaker was heated strongly during drying and the solid turned black? (Strong heating can oxidize Cu to black CuO.)
2. Why was the mixture of Fe and $CuSO_4$ heated? Suggest changes that might result if the mixture is not heated.

Part II. Data Analysis: Limiting Reactant

6. From your balanced equation, what do you expect a graph of mol Cu *vs.* mol $CuSO_4$ looks like?
7. Using the class data, plot a graph of mol Cu *vs.* mol $CuSO_4$. Does the graph look like your prediction?
8. Is there a direct relationship between moles Cu formed and moles $CuSO_4$ added?
9. Using the mole ratio from the balanced equation, calculate the moles of $CuSO_4$ that would react with the total moles of Fe used. Also calculate the theoretical number of moles of Cu that would be formed from this amount of Fe.
10. Using the mole ratio from the balanced equation, calculate the moles Fe that would react with the total moles $CuSO_4$ used. Also calculate the theoretical moles Cu that would be formed from this amount of $CuSO_4$.
11. Compare the theoretical (calculated) moles Cu with the moles Cu obtained in the reaction. The reactant ($CuSO_4$ or Fe) that gives the number of moles Cu closest to the experimental value is called the *limiting reactant* since this substance limits the yield of products. Calculation of moles of product(s) from the moles of limiting reactant agrees with experimental results. What is the limiting reactant?

    The other reactant is said to be present *in excess*. The theoretical moles of this reactant is the amount that will react with the moles of limiting reactant actually used in the reaction. Some of the reactant in excess remains “left over” when the reaction is complete. What reactant is present in excess?

Part II. Implications and Applications

3. When the reaction was complete, you recorded your observations. What observation indicates that $CuSO_4$ is in excess? How would your observations differ if Fe were the excess reactant?
4. Why is the graph of mol Cu *vs.* mol $CuSO_4$ different from the graph of mol Cu *vs.* mol Fe?
5. a. Calculate the mass of $CuSO_4$ reacted.

    [Alternate procedure: Calculate the volume of 0.50 M $CuSO_4$ solution reacted.]

    b. How would your experimental results differ if you used less than the quantity of $CuSO_4$ calculated in Question 5a?
6. Calculate the mass of $FeSO_4$ that could be obtained in your reaction.
7. In most laboratory chemical reactions or in chemical manufacturing processes, there is a limiting reactant and all other reactants are present in excess. Why do you think that is so?
8. If you were in charge of a chemical plant, suggest some practical considerations that would help you to decide which reactant should be the limiting reactant.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Stoichiometry of a Reaction and Limiting Reactant

Major Chemical Concept

The purpose of this activity is to collect data that leads to establishing the mole relationship between coefficients in the balanced equation, and to introduce the concept of limiting reactant.

Level

Analysis of the data collected to determine the mole relationship between coefficients of reactants and products (*Part I*) is appropriate for all levels. Extension of the data analysis to develop the concept of limiting reactant (*Part II*) may not be appropriate for general students with very weak mathematics skills but would be appropriate for general and honors students.

Expected Student Background

For *Part I*, students should be able to make gram-to-mole conversions. They should be able to plot and read graphs and understand the concept of ratios. For *Part II*, students will also need to understand how to use the stoichiometric ratio (mole ratio) to calculate moles of product from moles of reactant.

Time

Experimental data can be collected in one class period (50 min) if a drying oven is used to dry the Cu. If the Cu is allowed to dry overnight, weighing of Cu and analysis of data can be completed on the second day. (Be sure beakers are left where they will not be disturbed and they are labeled so students can identify their beaker. Do not put stick-on labels on the beakers—this would change the mass of the beaker and invalidate the experimental results. One way to label beakers is to place a piece of paper under the beaker indicating students' names in the group and the class period.)

Safety

Read *Safety Considerations* in the *Student Version*. Although $CuSO_4$ is not considered a dangerous substance, any spills (of solution or crystals) should be cleaned up immediately. The soluble product, $FeSO_4$, is also considered reasonably safe.

Materials (For 24 students working in pairs)

Nonconsumables

12 Beakers, 150- to 250-mL
12 Beakers, 50- to 100-mL
12 Graduated cylinders, 25-mL
Balance (mg accuracy recommended)
12 Burners or hot plates
If using burner: striker or other means of lighting burner, ring stands, rings, wire gauzes
Drying oven or heat lamps for drying Cu (optional)*

Consumables

Iron filings, 6.0 g
Copper(II) sulfate, $CuSO_4 \cdot 5H_2O$, 36 g or 500 mL 0.50 M $CuSO_4$
Distilled water, 1.0 L (optional)
Detergent solution, 10 mL
*Optional: You may choose to dry the precipitate with acetone.

Advance Preparation

All equipment should be readily available for student use. If you elect to have students use 0.500 M $CuSO_4$, the solution is prepared by dissolving 62.4 g $CuSO_4 \cdot 5H_2O$ in sufficient distilled water to give 500 mL solution. You may use a graduated cylinder for measuring the quantity of the solution. Be sure all solid is dissolved before diluting to final volume. Thoroughly mix the solution and transfer to a container labeled "0.50 M $CuSO_4$." Distilled water is preferable to tap water, but the results will not be significantly affected if tap water is used.

Depending upon the number of student groups, select masses of iron filings to be used by each group. Masses should be between 0.2 and 0.5 g. This will allow for a difference in the mol Fe used and consequently the mol Cu formed. Different masses of Fe should be assigned each group to give a meaningful graph of mol Cu *vs.* mol Fe. Assign students approximate masses, *e.g.*, 0.20 g or 0.45 g. The exact mass should be determined, but it does not need to be exactly 0.200 g or 0.450 g to show the mole relationship. (The quantity of $CuSO_4$ will be in excess as long as the mass of Fe is less than 0.67 g.)

Check with your local waste disposal official about proper disposal of the $CuSO_4$ waste. $CuSO_4$ should not be flushed down the drain. Metals more active than Cu will displace Cu from $CuSO_4$. Lead can be dissolved by $CuSO_4$ that remains in plumbing traps, causing water leaks in drain lines. (The word *plumbing* comes from plumbum, lead.) Some states enforce special disposal requirements for copper salts. Collect the decantate from the reaction for proper disposal. (It should not be necessary to retain the water from the washing of the Cu, only the first decantate should require special disposal procedures.) Even if your state currently has no special disposal requirements, it is good practice to make students aware that they should not indiscriminately flush waste down the drain or throw in the garbage.

Pre-Laboratory Discussion

During the pre-laboratory, review the mole concept and the relationship between molar mass and moles of a substance. Present the Law of Conservation of Mass.

Since students will heat the reaction mixture to accelerate reaction, explain that chemical reactions occur more rapidly at elevated temperatures. As illustrations, mention more rapid souring of milk when at room temperature than under refrigeration, and the burning of gasoline in air at high temperatures whereas, it can be kept at room temperature in the presence of air without reacting.

Explain that a solution of $CuSO_4$ is used rather than the solid because reactions occur more rapidly when the reactants come into close contact. $CuSO_4$ solution contains individual Cu^{2+} and SO_4^{2-} ions. Ions are free to move about in solution and can come into closer contact with the Fe than when confined to fixed positions in the crystal.

If nomenclature has been discussed, use this opportunity to review formulas and names of reactants and products. Do not write the reaction at this point since the purpose of *Part I* of the activity is to determine the mole relationship between Fe and Cu (and by implication $CuSO_4$ and $FeSO_4$).

Review safety precautions for the use of burners if they are used for heating the reaction mixture. Generally the reaction is complete when the reaction mixture begins to boil. Caution students not to allow the solution to boil or they may lose some product, particularly if bumping occurs. (*Bumping* results from local "hot spots" in the beaker where uneven and eruptive boiling occurs. Bumping is dangerous since the hot liquid can be spewed for some distance. Beakers have been known to "jump off the ring stand" during bumping!)

Explain that each group will follow the same procedure but will use a different quantity of Fe and that all data will be combined for a class analysis of the data. It is important to accurately record the quantities of chemicals used. Problems sometimes arise when iron filings become coated with copper and thus do not fully react.

If you have not discussed solutions and plan to have students use a $CuSO_4$ solution, explain briefly the meaning of molarity (moles $CuSO_4$ per liter of solution) and why molarity is a convenient way of expressing the solution's composition (easy to determine moles $CuSO_4$ by measuring the solution volume). You may prefer to lead the entire class in a calculation of moles $CuSO_4$ used in the reaction. A simple way to explain this would be to say: "If 1 L or 1000 mL contains 0.500 mol $CuSO_4$, how much $CuSO_4$ would 100 mL contain?

100 mL = $\frac{1}{10}$ of 1000 mL, thus $\frac{1}{10}$ of 0.500 mol = 0.0500 mol $CuSO_4$.

If 100 mL contains 0.0500 mol $CuSO_4$, how much $CuSO_4$ would 25 mL contain?

25 mL = $\frac{1}{4}$ of 100 mL, thus $\frac{1}{4}$ of 0.0500 mol = 0.0125 mol $CuSO_4$."

You might tell students how many moles (or grams) of $CuSO_4$ are in 25 mL of the $CuSO_4$ solution (0.0125 mol or 2.00 g $CuSO_4$). (*NOTE: If your students weigh their own* $CuSO_4 \cdot 5H_2O$ *and dissolve it, about 3 g is required because the inclusion of the water of hydration in* $CuSO_4 \cdot 5H_2O$ *increases the molar mass from 159.7 g/mol to 249.7 g/mol.*)

If a drying oven is used, the sample should not be dried overnight as Cu oxidation may occur at the elevated temperature.

Teacher-Student Interaction

Move through the groups insuring proper safety procedures. Try not to answer many questions. Answer a question with a question when possible to facilitate developing thinking skills. Possible questions are:

1. What does the blue decantate indicate? *[Since* $CuSO_4$ *is used in excess, when the reaction is complete there will be some* $CuSO_4$ *remaining. Be sure students note the color of the decantate. This is one of the follow-up questions.]*
2. Why is heating continued after iron filings are added? *[To speed up the reaction.]*
3. Why is a solution of $CuSO_4$ used rather than just mixing crystals with iron filings? *[This is discussed in the* Pre-Laboratory Discussion *section.]*
4. Why does the activity call for iron filings instead of iron nails or other large pieces of iron? *[Because reaction is more rapid for finely divided solid.]*
5. Why is the product washed? *[To remove impurities,* $FeSO_4$ *and excess* $CuSO_4$*, from the Cu.]*

If students weigh $CuSO_4 \cdot 5H_2O$ and dissolve it, they will find that it does not dissolve readily.

6. How can you speed up the dissolving process for $CuSO_4 \cdot 5H_2O$? *[Stirring, grinding the crystals, heating the solution. Stirring and heating are methods used in this activity.]*
7. Why might grinding the crystals in a mortar and pestle not be a good idea? *[Unless the crystals are ground before weighing, the mass weighed and the mass dissolved might not be the same since some crystals might be lost (left on mortar and pestle) during crushing and not transferred to the solution.]*

Anticipated Student Results

Sample Data Table

1.	Mass of 50-mL beaker	26.292 g
2.	Mass of 50-mL beaker and iron filings	26.603 g
3.	Mass of iron filings	0.311 g
4.	Mass of 150-mL beaker	98.325 g
5.	Mass of 150-mL beaker and $CuSO_4{\cdot}5H_2O$	101.367 g
6.	Mass of $CuSO_4{\cdot}5H_2O$	3.042 g
7.	Mass of 150-mL beaker and product (Cu)	98.673 g
8.	Mass of dry product (Cu)	0.348 g

Part I. Data Analysis: Stoichiometry of the Reaction

1. mol Fe: $0.311 \text{ g Fe} \times \dfrac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 5.57 \times 10^{-3} \text{ mol Fe}$

2. mol $CuSO_4$:$3.042 \text{ g } CuSO_4{\cdot}5H_2O \times \dfrac{1 \text{ mol } CuSO_4{\cdot}5H_2O}{249.70 \text{ g } CuSO_4{\cdot}5H_2O} =$

$1.218 \times 10^{-2} \text{ mol } CuSO_4{\cdot}5H_2O$

$1.218 \times 10^{-2} \text{ mol } CuSO_4{\cdot}5H_2O = 1.218 \times 10^{-2} \text{ mol } CuSO_4$

If the alternate procedure is used, mol $CuSO_4$:

$$0.0252 \text{ L solution} \times \frac{0.500 \text{ mol } CuSO_4}{1 \text{ L solution}} = 1.26 \times 10^{-2} \text{ mol } CuSO_4$$

3. $0.348 \text{ g Cu} \times \dfrac{1 \text{ mol Cu}}{63.54 \text{ g Cu}} = 5.48 \times 10^{-3} \text{ mol Cu}$
4. (Graph)
5. Slope should be approximately 1, this is, 1 mol Cu : 1 mol Fe.

Balanced equation: $Fe(s) + CuSO_4(aq) \longrightarrow Cu(s) + FeSO_4(aq)$

Part II. Data Analysis: Limiting Reactant

6. Graph predicted is a straight line with slope 1 (1 mol Cu/1 mol $CuSO_4$).
7. Graph does not look like prediction. Line is almost vertical.
8. There is no direct relationship between moles Cu formed and moles $CuSO_4$. The moles Cu vary even though moles $CuSO_4$ are essentially constant for all groups.
9. mol $CuSO_4$ and mol Cu (“theoretical if all Fe is used”):

$$5.57 \times 10^{-3} \text{ mol Fe} \times \frac{1 \text{ mol } CuSO_4}{1 \text{ mol Fe}} = 5.57 \times 10^{-3} \text{ mol } CuSO_4$$

$$5.57 \times 10^{-3} \text{ mol Fe} \times \frac{1 \text{ mol Cu}}{1 \text{ mol Fe}} = 5.57 \times 10^{-3} \text{ mol Cu}$$

10. mol Fe and mol Cu (“theoretical if all $CuSO_4$ is used”):

$$1.218 \times 10^{-2} \text{ mol } CuSO_4 \times \frac{1 \text{ mol Fe}}{1 \text{ mol } CuSO_4} = 1.218 \times 10^{-2} \text{ mol Fe}$$

$$1.218 \times 10^{-2} \text{ mol } CuSO_4 \times \frac{1 \text{ mol Cu}}{1 \text{ mol } CuSO_4} = 1.218 \times 10^{-2} \text{ mol Cu}$$

11. Fe is the limiting reactant. $CuSO_4$ is present in excess.

A table such as the following may be useful in comparing the experimental and theoretical results.

Comparison of Theoretical and Experimental Yields of Cu

	Experimental	**Theoretical (If all Fe is used)**	**Theoretical (If all $CuSO_4$ is used)**
mol Fe	5.57×10^{-3} mol	5.57×10^{-3} mol	1.218×10^{-2} mol
mol $CuSO_4$	1.218×10^{-2} mol	5.57×10^{-3} mol	1.218×10^{-2} mol
mol Cu	5.48×10^{-3} mol	5.57×10^{-3} mol	1.218×10^{-2} mol

Answers to Implications and Applications

Part I. Stoichiometry of the Reaction

1. a. No effect. $CuSO_4$ is in excess with 3 g. Adding more would not affect the yield of Cu.

   b. Experimental mol Cu would be less because Cu would be poured off with decantate and not recovered and weighed.

   c. Experimental mol Cu would be more because water added to the mass of Cu. The total mass of Cu and water was used to calculate mol Cu.

   d. Experimental mol Cu would be more because Cu was oxidized to CuO, which would have a greater mass than Cu. Since the total mass (Cu and combined O) was used to calculate mol Cu, the apparent mol Cu would be greater.

2. The mixture was heated to speed up the reaction. If the reaction mixture were not heated, the reaction would take longer. If the reaction did not go to completion before decanting, some Fe might remain unreacted and be weighed as Cu, causing an error in mol Cu.

Part II. Limiting Reactant

3. The blue color of the solution indicates $CuSO_4$ is present after the reaction is complete. If Fe were the reactant in excess, the decantate would not be blue. (Also Fe would be mixed with Cu and the mass of the remaining solid would not be the mass of Cu. Do not emphasize this point, however, unless you want to discuss how the activity could be modified to determine the mass of Cu. One possibility is to dissolve the excess Fe in HCl to recover Cu, which is insoluble in HCl.)

4. Since $CuSO_4$ is in excess, there is not a mole relationship between mol $CuSO_4$ added and mol Cu formed. The mole relationship exists only for the limiting reactant (Fe).

5. a. Mass $CuSO_4{\cdot}5H_2O$ (volume 0.500 M $CuSO_4$):
   1 mol $CuSO_4$ = 1 mol $CuSO_4{\cdot}5H_2O$

$$5.57 \times 10^{-3}\ \text{mol CuSO}_4{\cdot}5\text{H}_2\text{O} \times \frac{249.7\ \text{g CuSO}_4{\cdot}5\text{H}_2\text{O}}{1\ \text{mol CuSO}_4{\cdot}5\text{H}_2\text{O}} = 1.39\ \text{g CuSO}_4{\cdot}5\text{H}_2\text{O}$$

$$\left[5.57 \times 10^{-3}\ \text{mol CuSO}_4 \times \frac{1\ \text{L solution}}{0.500\ \text{mol CuSO}_4} = 0.0111\ \text{L} = 11.1\ \text{mL solution}\right]$$

   b. $CuSO_4$ would be the limiting reactant; the experimental mass (and amount in moles) Cu recovered would be less.

6. Mass $FeSO_4$:

$$5.57 \times 10^{-3}\ \text{mol Fe} \times \frac{1\ \text{mol FeSO}_4}{1\ \text{mol Fe}} \times \frac{152.9\ \text{g FeSO}_4}{1\ \text{mol FeSO}_4} = 0.852\ \text{g FeSO}_4$$

7. Many students recognize that it is difficult to measure exactly equivalent mole amounts of all reactants. (Another reason is that adding an excess of one reactant insures that all of the other reactant is used in the reaction.)

8. Practical considerations include: Which reactant is easier to measure? Which reactant is less expensive?

Post-Laboratory Discussion

Reinforce the significance of moles in chemical reactions. Be sure students understand that coefficients in an equation relate to moles (not grams) of reactants and products. Discuss the concept of limiting reactant. Analogies are particularly helpful (see *Metaphors and Analogies*). Review student responses to the questions in *Implications and Applications*. Item 1 in *Possible Extensions* could be used during post-laboratory discussion.

Possible Extensions

1. An alternative treatment of data is to have students calculate the ratio mol Cu/mol Fe and plot the number of students having a given mole ratio (on the y-axis) *vs.* the mole ratio (on the x-axis). Mole ratios should be rounded to ± 0.1; *i.e.*, plot ratios of 0.7, 0.8, 0.9, 1.0, 1.1, *etc.* It should be clear that most data (the middle two-thirds) fall very close to 1.0. The range of the middle two-thirds of the mol Cu/mol Fe data can be noted and compared with uncertainty in the measurements.
2. Repeat the activity with other metals such as Al, Mg, Zn. Mole ratios can be determined and equations written. (Adjustments in the quantity of metal may have to be made to insure that $CuSO_4$ is present in excess.) In Activity S5 “What Is the Copper Formula?” in *Laboratory Assessment Builds Success* the percent copper in $CuCl_2$ is determined by reaction of Al with $CuCl_2$. The activity could be adapted to illustrate stoichiometry if the data is treated as in this activity.
3. If Mg or Zn is used and students are told that reaction stoichiometry is the same as for the reaction of $CuSO_4$ with Fe, the molar mass of the metal can be determined from data collected (mass of metal reacted and mass of Cu formed).
4. *Activity 1: Limiting Reagent for a Chemical Reaction in Solution* in the *Solutions* module illustrates solution stoichiometry.

Assessing Laboratory Learning

1. **Laboratory Practical.** Give students sample data (mass or mol Fe and $CuSO_4$) and ask them to determine which is the limiting reagent and the theoretical yield of Cu. Give them data for mol of unknown metal (M) and mol $CuSO_4$, and ask them to plot the data to determine the reaction stoichiometry. (Alternately, a graph of mol unknown metal (M) *vs.* mol Fe could be given and students asked to determine the stoichiometry.)
2. **Laboratory Report.** Although the data for the entire class are used for the graph to determine the reaction stoichiometry, it is worthwhile to have students plot their own graph as part of the activity report. Explain that graphing is important in analysis of chemical data and that they should become familiar with techniques for plotting and reading graphs. Also important is the idea that use can be made of the same data for different purposes: to determine the stoichiometry of the reaction and to demonstrate which reactant is the limiting reactant. Because the data analysis is somewhat involved, you may divide the laboratory report into two parts or use an abbreviated format. An abbreviated format might consist of tables, graphs and brief discussion of the meaning of the graphs and tables with calculations in an appendix.
3. **Written Examination.** Some possible items:
   a. An unknown metal reacts with an excess of $CuSO_4$ to form Cu. The following data are determined for reaction of the unknown metal with $CuSO_4$:

mol metal(M) reacted	mol Cu formed
2.80×10^{-3}	4.20×10^{-3}
3.70×10^{-3}	5.55×10^{-3}
4.11×10^{-3}	6.17×10^{-3}
5.52×10^{-3}	8.28×10^{-3}

   Plot mol Cu *vs.* mol M and find the coefficients of M (the unknown metal) and Cu in the balanced equation. *[3 mol Cu : 2 mol M]*

   b. When 0.859 g Fe is reacted with 3.47 g $CuSO_4$, what is the theoretical yield of Cu? *[0.979 g Cu—Fe is the limiting reactant.]*

References

Cotton, F. A., Darlington, C. L., and Lynch, L. D. (1980). *Chemistry: An investigative approach*. Boston, MA: Houghton Mifflin.

Gardner, M., Smith, P., and Heikkinen, H. (1990). *Laboratory assessment builds success*. Berkeley, CA: Lawrence Hall of Science, Institute for Chemical Education.

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Percent Yield of a Chemical Reaction

Introduction

The yield for a chemical reaction, the relationship between that yield, and the balanced equation representing the reaction are very important to chemists. Not only does a laboratory chemistry investigation require an understanding of such concepts, but the chemical industry also depends on percent yield principle in the manufacture of chemical products.

The reaction in this activity uses sodium bicarbonate and hydrochloric acid. The equation for this reaction shows the mole relationships for the reaction. The yield of NaCl will be determined and compared with the theoretical yield.

$$HCl(aq) + NaHCO_3(s) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

Purpose

To illustrate the relationship between the balanced equation and the amounts of substances involved in a reaction either as reactants or products. To determine the percent yield of a reaction.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Exercise caution when adding the HCl solution to sodium bicarbonate to avoid splashing during reaction. When drying the product, heat slowly and cautiously to avoid spattering. HCl can cause severe burns and can irritate the respiratory tract.
3. If burners are used to evaporate the solution, exercise caution in lighting the burner. Watch out for loose hair. Hair ignites easily especially if hair spray is used.
4. Always exercise caution when handling glassware, especially if the glassware is hot. Hot and cold glass look the same!

Procedure

1. Weigh a clean dry 250-mL beaker.
2. Add approximately 2.5 g $NaHCO_3$ to the beaker.
3. Accurately weigh the beaker and $NaHCO_3$.
4. Slowly add the HCl solution (about 10-15 mL) and observe. You may be told to use a dropper to add the acid or just pour very slowly.
5. When the reaction has apparently stopped, add a few more milliliters of the HCl solution to be sure all $NaHCO_3$ has reacted.
6. Evaporate the solution to recover the product, NaCl. Your teacher will tell you how to evaporate the resulting solution to obtain dry NaCl. You may use a drying oven or a burner. If a burner is used, heat slowly and carefully to avoid spattering. Heat in a hood. Be careful not to breathe fumes.
7. Thoroughly wash your hands before leaving the laboratory.

Observations

Make observations about substance color, nature of the reaction, *etc.*, as you perform the laboratory.

Data Analysis

1. Prepare a data chart as shown:

   Mass of empty beaker ____________________

   Mass of beaker and $NaHCO_3$ ____________________

   Mass of $NaHCO_3$ ____________________

   Mass of beaker and NaCl after evaporating ____________________

   Mass of NaCl recovered ____________________

2. From grams $NaHCO_3$ used, calculate mol $NaHCO_3$ used and mass NaCl formed (theoretical yield) using the balanced equation.
3. Using the mass NaCl recovered, calculate mol NaCl recovered.
4. Calculate the percent yield for the reaction from the actual yield (g NaCl recovered) and the theoretical yield (g NaCl calculated in Question 2).

$$\%\text{ Yield} = \frac{\text{Actual yield recovered}}{\text{Theoretical yield}} \times 100$$

5. Share class data. Suggest reasons why the percent yield does not equal 100%. (It may be greater than 100%!)

Implications and Applications

1. Why can the percent yield *theoretically* not be greater than 100%? Use the Law of Conservation of Mass to support your answer.
2. Why is it necessary to add sufficient HCl to ensure the complete reaction of $NaHCO_3$?
3. What could happen to the percent yield in each of the following situations:
   a. The mass of $NaHCO_3$ used is recorded as 2.345 g; it is actually 2.545 g?
   b. The mass of NaCl recovered is recorded as 1.786 g; it is actually 1.886 g?
   c. Some distilled water is accidentally added to the beaker containing NaCl before the solution is evaporated?
4. What effect would using a more dilute HCl solution have on the percent yield?

*5. Suppose $NaHCO_3$ used in this activity is contaminated with NaCl. How would this affect the percent yield? (See comments in *Teacher Notes*.)

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Percent Yield of a Chemical Reaction

Major Chemical Concept

Data collection to support the stoichiometric relationships based on the balanced equation, and the introduction of the percent yield concept are the main goals of this activity. The emphasis is on the relationship between the balanced equation and the actual amounts of substances used and produced in this chemical reaction. The concept of limiting reactant while not stressed in this activity is implied in the part where students are instructed to add additional HCl after the reaction had apparently stopped. One of the student questions relates to this. It is *not* too early to introduce the idea of excess and limiting reactants. Comparison of class data for this activity shows that percent yields are independent of the amounts of reactants.

One major weakness of this activity is all mole ratios are 1:1, which may tend to lead to a careless assumption that moles of reactant are always equal to moles of product. This laboratory should be followed by a thorough discussion of ratios other than 1:1, as well as those that are 1:1.

Level

All levels of introductory chemistry should be able to complete this activity, with varying amounts of teacher input and guidance. (Question 5 in *Implications and Applications* may be too difficult for general-level students.)

Expected Student Background

Students should be able to make mole conversions (grams to moles) and to compute percentage. Ratio and proportion concepts will be needed to connect actual molar amounts used in the laboratory and mole ratios according to the balanced equation.

Time

This activity can be completed easily in less than 50 min if the proper pre-laboratory is provided and a drying oven is used to evaporate the solution containing the NaCl. If burners are used to evaporate the solutions, finish the reaction one day and complete the evaporation the next. If a drying oven is used, weighing the next day takes very little time. This allows ample class time to collect class data and make calculations.

Safety

Read the *Safety Considerations* in the *Student Version*.

Materials (For 24 students working in pairs)

Nonconsumables

12 Beakers, 250-mL
Balance (mg accuracy recommended)
12 Burners or drying oven
If using burner: striker or other means of lighting burner, ring stands, rings, wire gauzes
If drying oven is used: aluminum foil

Consumables

Sodium bicarbonate, $NaHCO_3$, 24-36 g
3 M Hydrochloric acid, HCl, 250 mL (62.5 mL conc. HCl diluted to 250 mL)

Advance Preparation

Prepare 3 M HCl and place in dropper bottles for each laboratory team.

Check all equipment ahead to be sure all is in good working condition. Having equipment in plain view at laboratory stations will save much time! *WARNING: If a drying oven is used, cover the beaker with aluminum foil.* HCl vapors will react with the walls of many drying ovens causing rapid oxidation (rusting). Any aluminum chloride formed by the reaction of HCl with aluminum foil clings to the aluminum foil and will not affect experimental results.

Decide whether to assign Question 5 in *Implications and Applications* or use it in the post-laboratory discussion.

Pre-Laboratory Discussion

The purpose of this activity is to study stoichiometry, not to predict reaction products. Review stoichiometric relationships and the calculations required in this activity. During the pre-laboratory, consider demonstrating the addition of HCl solution to $NaHCO_3$. Have students feel the beaker before and after adding HCl to $NaHCO_3$. It should become noticeably cooler. Proper laboratory skills, such as lighting burners, should be stressed. Thorough safety instructions must be given. If you choose to omit Question 5 in *Implications and Applications*, inform the students.

Possible pre-laboratory discussion questions are:

1. Why is additional HCl solution added? *[To insure the reaction is complete.]*
2. Which product might be recovered easily for purposes of measuring the amount? *[Consider the physical states of the products. It should be obvious that carbon dioxide gas would be difficult but not impossible to recover. It is not feasible to try to determine how much water is made since the amount is so small compared to the amount of water in the HCl solution. $NaCl$ would be the easiest product to recover.]*
3. What data should be collected? *[See data chart in* Student Version.*]*
4. How could you use collected data to determine the stoichiometry of the reaction? *[See* Activity 1: Stoichiometry of a Reaction and Limiting Reactant.*]*
5. How could you use the data to determine percent yield? *[This is described in Questions 2-4 in* Data Analysis *of* Student Version. *Be sure to include in your discussion the need to collect class data for comparison of results.]*

Teacher-Student Interaction

Move among groups to insure proper safety procedures. Try not to answer many questions. Answer a question with a question when possible to facilitate developing thinking skills. Possible questions are:

1. Did you notice a temperature change in the beaker during the reaction? *[Reaction is endothermic—beaker gets cool. Although most neutralization reactions are exothermic, this reaction is endothermic. Students may have observed that the glass gets cold when Alka-Seltzer (principally citric acid and* $NaHCO_3$*) is dissolved in water. You might wish to point out that this activity is essentially the same reaction (Acid + Bicarbonate).]*
2. What evidence do you observe for a chemical reaction? *[Fizzing, beaker gets cool.]*
3. What causes the fizzing? *[*$CO_2(g)$*]*

This laboratory can be very quickly completed leaving time for the discussion about observations.

Anticipated Student Results

An actual data chart from a student's laboratory report:

Mass of empty beaker	93.650 g
Mass of beaker and $NaHCO_3$	95.151 g
Mass of $NaHCO_3$	1.501 g
Mass of beaker and NaCl after evaporating	94.691 g
Mass of NaCl	1.041 g

Answers to Data Analysis

1. An actual data chart from a student's laboratory report is shown in *Anticipated Student Results*.
2. $$1.501\text{ g NaHCO}_3 \times \frac{1\text{ mol NaHCO}_3}{84.01\text{ g NaHCO}_3} = 1.787 \times 10^{-2}\text{ mol NaHCO}_3$$
   $$1.501\text{ g NaHCO}_3 \times \frac{1\text{ mol NaHCO}_3}{84.01\text{ g NaHCO}_3} \times \frac{1\text{ mol NaCl}}{1\text{ mol NaHCO}_3} \times \frac{58.44\text{ g NaCl}}{1\text{ mol NaCl}} = 1.044\text{ g NaCl}$$
3. $$1.041\text{ g NaCl} \times \frac{1\text{ mol NaCl}}{58.44\text{ g NaCl}} = 1.781 \times 10^{-2}\text{ mol NaCl}$$
4. $$\frac{1.041\text{ g NaCl}}{1.044\text{ g NaCl}} \times 100\% = 99.7\text{ \% yield}$$
5. Spattering during the reaction or during evaporation would give low results. Water (sample is not dry) or other impurities could cause high results. (Further drying should improve results if the sample is wet.) Incomplete reaction of $NaHCO_3$ will also give high results. Incorrect masses could result in either high or low results.

Answers to Implications and Applications

1. The percent yield theoretically should not be greater than 100% because the Law of Conservation of Mass states that matter is not created or destroyed in ordinary chemical reactions. The concept of a balanced equation is based on the Law of Conservation of Mass.
2. Theoretical yield is based on the mass of $NaHCO_3$ added.
3. a. Since mass of $NaHCO_3$ is too small, the theoretical yield of NaCl will be too small. The calculated percent yield will be too high (probably greater than 100%).
   b. The actual yield of NaCl is too small. The calculated percent yield will be too low.
   c. There would be no effect on the percent yield since the water is evaporated.
4. None. It would simply require more HCl solution to complete the reaction.
5. If $NaHCO_3$ is contaminated with NaCl, the assumed mass of $NaHCO_3$ would be too large, making the theoretical yield of NaCl too large. Thus the calculated percent yield will be too low. In discussing this with students, select masses for $NaHCO_3$ and NaCl (*e.g.*, 1.00 g NaCl and 1.00 g $NaHCO_3$) and calculate "theoretical yield" (assuming 2.00 g $NaHCO_3$) and "actual yield" (assuming 1.00 g NaCl and 1.00 g $NaHCO_3$).

Post-Laboratory Discussion

Post-laboratory discussion should start by collecting class data. Collect the data for moles $NaHCO_3$ used, the moles NaCl produced, and the percent yield for each group. The class should be able to determine that percent yield is independent of how much reactant is used. Use moles NaCl collected compared to moles NaCl predicted as well as grams collected compared to grams predicted to calculate the percent yield. This is another way of reinforcing the concept that mole is an amount. This activity is fairly consistent in giving very close to 100% yields. Discussing how the mole ratio from the balanced equation is used in this laboratory is essential. Ask such questions as “Does the amount of $NaHCO_3$ used affect the amount of NaCl produced?” and “Why does the mass of NaCl characteristically seem to be less than the mass of $NaHCO_3$ used if the mole ratio is 1:1?” This leads to the discussion of different molar masses for compounds.

Propose the idea that the students are chemists in a chemical manufacturing plant in charge of deciding how much of each reactant they need to have on site for the production of certain chemicals. How would percent yield for their product enter into the decision about amounts of reactants to order?

Extensions

1. An important part of chemical education is the continued integration of all disciplines in the learning process. After collecting the data for moles $NaHCO_3$ and moles NaCl involved in this reaction, have students plot a graph of mol $NaHCO_3$ *vs.* mol NaCl. The slope of the line is close to one as expected from the balanced equation. This discussion should lead to the question, if the ratio is not 1:1 for a particular comparison being made, what happens to the slope? Can the slope of such a line be used to determine mole ratios of reactions?
2. If you have discussed gases, you might ask students to suggest ways to determine how much CO_2 is generated in the reaction. The most frequent answer is to collect the gas by water displacement. Discuss why this would not be practical for CO_2. *[CO_2 dissolves in water as evidenced by its use to “carbonate” beverages.]* Students could research methods for determining CO_2.
3. Students need to be aware that data is not valid unless statistically evaluated. Simple statistical analysis could involve calculating the average and standard deviation of class data. Students do not need to understand the theory behind the formulas used to derive standard deviation. Choose one of the available computer software packages for the calculations. Connect the average and standard deviation to the normal curve to give students an appreciation for the ranges that exist for measured quantities. (The mathematics and science departments could team-teach here.)

Assessing Laboratory Learning

1. **Laboratory practical.** Provide students with sample data from a different laboratory activity and have them make the appropriate calculations to determine percent yield. Try to use data that requires a different approach to the same concepts.

   $$MgSO_4(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + Na_2SO_4(aq)$$

   500 mL of 0.429 M solution $MgSO_4$ was treated with 0.500 M NaOH.

   a. How much NaOH was necessary to precipitate the $Mg(OH)_2$ completely? *[858 mL]*
   b. Assuming 100% yield, how many g of $Mg(OH)_2$ can be expected to form? *[12.5 g]*
   c. The actual yield for this reaction is 10.4 g. What is the percent yield?*[83.2%]*
   d. Can you think of a reason why the percent yield is so low? *[$Mg(OH)_2$, although only slightly soluble in water, has a fairly large solubility product constant (see* Solubility and Precipitation *module). If a filtration method is used to separate the precipitated $Mg(OH)_2$ from the supernatant liquid, an appreciable amount of $Mg(OH)_2$ will remain in solution, contributing to the low yield.]*

2. **Laboratory report.** One variation on the individual laboratory report is a group laboratory report. With the latter, the group consists of perhaps 4 or 5 pairs of students who worked separately. After completing the laboratory, all pairs combine data and analyze the outcome. Requirements can be more stringent. The laboratory report must be typed, double spaced,*etc.* The grading should be done on a very competitive basis. It is better to have another competent and trusted teacher evaluate the reports on the basis of information presentation. This makes the activity more realistic because in real-world situations groups of individuals combine efforts for a final product, often with outside evaluation.
3. **Examination questions.** Some possible items:
   a. Three samples of a given metallic ore are weighed. The masses are 12.45 g, 34.65 g, and 78.90 g, respectively. Analysis of the ores reveals that the samples contain 7.22 g, 21.14 g, and 44.97 g of copper, respectively. Based on these data, if a metallurgist wished to reclaim the copper from 125 kg of this ore, and knew that the process required 4 mol of O_2 gas for every 2 mol of copper to be recovered, how many mol O_2 gas would be needed? *[The % Cu in the ore is 58.7% (average of 3 analyses that give 57.0%, 61.01% and 57.00%, respectively), 125 kg ore would contain 73.4 kg Cu (1.15 x 10^3 mol Cu) and would require 2.30 x 10^3 mol O_2.]*
   b. A reaction has a percent yield of 36.7%. If the reaction involves the production of CO_2 gas at this percent yield, how many moles of Reactant A would be needed to give 88.0 g of CO_2 if the mole ratio is 3 mol A : 1 mol CO_2?

$$88.0\ g\ CO_2\ x\ \frac{1\ mol\ CO_2}{44.0\ g\ CO_2}\ x\ \frac{3\ mol\ A}{1\ mol\ CO_2} = 6.00\ mol\ A$$

$$36.7\%\ X = 6.00\ mol\ A$$

$$X = \frac{6.00\ mol}{0.367} = 16.3\ mol\ A$$

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Balancing Chemical Equations

Purpose

To illustrate the relationship of the Law of Conservation of Mass to balanced chemical equations.

Materials

Gumdrops
Toothpicks

Safety

No special precautions are needed. Students should not eat the candy, particularly if this is done in the laboratory.

Directions

Using gum drops (or jelly beans) for atoms and toothpicks to hold atoms together, prepare models of reactant compounds (“reactant molecules”). Different colors of gum drops should be used to represent different elements. Several “molecules” of each reactant should be made. The atoms are then rearranged to form the product “molecules.” All gum drops from any “molecule” that reacts must be used. None may be lost (eaten?!) and “extra atoms” not in the reactant “molecules” may not be used to make products. When the “product molecules” have been made, the reaction is balanced. The balanced equation is written for the “reaction.”

The purpose is to emphasize the Law of Conservation of Matter in balancing equations. The arrangement of atoms in the molecules is unimportant. The purpose is to illustrate the Law of Conservation of Mass in balancing equations, not molecular geometry. Students should be given several *unbalanced* reactions for which to make models and write a balanced equation. Examples of appropriate reactions (unbalanced) are:

$$\begin{aligned}
H_2 + O_2 &\rightarrow H_2O \\
NH_3 &\rightarrow N_2 + H_2 \\
H_2S + SO_2 &\rightarrow H_2O + S \\
P + Cl_2 &\rightarrow PCl_5 \text{ (or } PCl_3\text{)} \\
CH_4 + O_2 &\rightarrow CO_2 + H_2O
\end{aligned}$$

(or combustion reaction for other hydrocarbons, such as C_2H_6, C_2H_4, C_3H_8)

Some more difficult reactions to balance:

$$\begin{aligned}
NO_2 + O_2 &\rightarrow N_2O_5 \\
NH_3 + O_2 &\rightarrow NO + H_2O \\
PBr_3 + H_2O &\rightarrow H_3PO_3 + HBr \\
PCl_5 + H_2O &\rightarrow H_3PO_4 + HCl \\
SCl_4 + H_2O &\rightarrow H_2SO_3 + HCl
\end{aligned}$$

Demonstration 2: Electrolysis of Water

Purpose

To illustrate reaction stoichiometry involving gaseous products.

Materials

Electrolysis apparatus (The *Electrochemistry* module has information on electrolysis and construction of electrolysis apparatus.)
Water
6 M Sulfuric acid, H_2SO_4, 10 mL
Wooden splint

Safety

Do not allow the reaction to proceed long enough to collect large quantities of hydrogen gas. Hydrogen gas can be explosive if ignited.

Directions

Electrolysis of water will yield two volumes of hydrogen for every volume of oxygen. Since water will not conduct an electric current, a small amount of an electrolyte that will not be oxidized or reduced as easily as water (usually sulfuric acid) is added to the water to provide ions to carry the charge. The identity of the two gases can be determined by showing that a glowing splint will burn when thrust into the oxygen, but not the hydrogen.

Demonstration 3: Limiting Reactant

Purpose

To illustrate the limiting reactant concept.

Materials

Household vinegar (5% acetic acid, $HC_2H_3O_2$), 100 mL
Sodium bicarbonate, $NaHCO_3$ (baking soda), 5 g
Funnel
6 Party balloons
6 Test-tubes, 18- x 150-mm

Safety

No special safety precautions are needed

Directions

1. Add 10.0 mL of household vinegar to each of six 18- x 150-mm test-tubes. Place these test-tubes in a test-tube rack.
2. Obtain six small round party balloons, the same size and preferably the same color.
3. Weigh one of the balloons. Using a funnel add 0.18 g of household baking soda. Be sure that the baking soda goes to the bottom of the balloon and none remains near the opening.
4. Repeat Step 3 for the remaining five balloons containing 0.35 g, 0.52 g, 0.70 g, 1.00 g and 1.70 g, respectively. (The balloons may be prepared ahead of time for demonstration the following day.)
5. Attach each of the six balloons to the test-tubes prepared in Step 1, taking care not to mix the contents of the balloons and test-tubes (see Figure 1).
6. After the balloons are securely attached to the test-tubes, lift the balloons one at a time and allow mixing of the contents of the balloon and the test-tubes. Enlist student assistants for this part of the demonstration, especially to secure the balloon to the test-tubes during inflation.

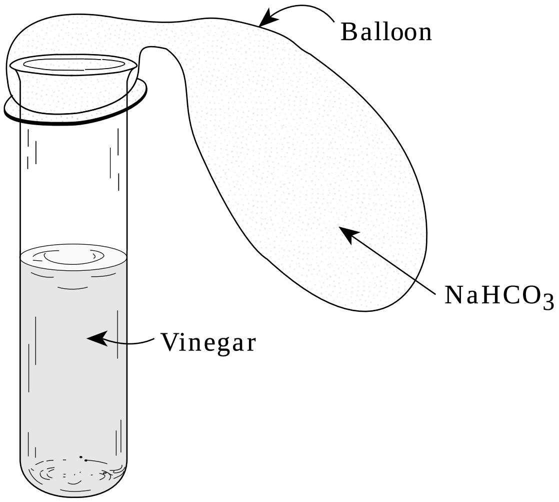


Figure 1. Diagram for limiting reactant demonstration.

Observations: Balloons should inflate with the gaseous product of this reaction. Balloons attached to Tubes 1 though 4 should gradually increase in volume. Balloons attached to Tubes 4 through 6 should be approximately the same size, even though more baking soda has been added to these balloons.

1. Which reactant was limiting in test-tube . . .

   1? 2? 3?

   4? 5? 6?

   [Tubes 1-3: $NaHCO_3$; Tubes 5-6 $HC_2H_3O_2$; Tube 4 has approximately stoichiometric proportions of the reactants.]

   Reaction occurring is: $2\ H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$

Calculations:

2. Household vinegar is 5% acetic acid, $HC_2H_3O_2$ (MM = 60 g/mol). If the density of household vinegar is approximately 1.00 g/mL, how many moles of acetic acid are contained in 10 mL of household vinegar? *[8.3×10^{-3} mol]*
3. Find the number of moles of sodium bicarbonate, $NaHCO_3$ (MM = 84 g/mol; baking soda) contained in each of the balloons:

Balloon 1: 0.18 g ____ *[2.1×10^{-3} mol]* Balloon 2: 0.35 g ____ *[4.2×10^{-3} mol]*
Balloon 3: 0.52 g ____ *[6.2×10^{-3} mol]* Balloon 4: 0.70 g ____ *[8.3×10^{-3} mol]*
Balloon 5: 1.00 g ____ *[1.2×10^{-2} mol]* Balloon 6: 1.70 g ____ *[2.0×10^{-2} mol]*

4. Consider answers to calculation Questions 2 and 3 and your observations about limiting reactant in Question 1. Do they agree? Explain the relationship of the calculations to your observations.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. How is the Law of Conservation of Mass related to a balanced equation? *[In a balanced equation the mass of reactants equals the mass of products.]*
2. How are moles related to reaction stoichiometry?*[A balanced equation relates moles of products to moles of reactants. Since moles is related to mass of a substance, masses of reactants and products can be related.]*
3. What is the relationship between limiting reactant and theoretical yield? *[Since the limiting reactant is present in smallest stoichiometric amount, it determines theoretical yield.]*
4. How does percent yield support the idea that actual yield cannot be more than theoretical yield? (How does this relate to the Law of Conservation of Mass?) *[Because mass is conserved in all chemical reactions, percent yield can not be greater than 100%. This implies that the theoretical yield is the maximum yield possible; thus, actual yield must be no more than the yield calculated from the stoichiometric relationship. (If actual yield appears greater, it is because of experimental error or because impurities are present.)]*
5. Why does the calculation of the theoretical yield imply that the reaction goes to completion? *[If the reaction does not go to completion, reactants will remain, and the actual yield will be less than the calculated yield.]*

Counterintuitive Examples and Discrepant Events

The puzzling nature of chemical reactions is illustrated by bubbling carbon dioxide into a calcium hydroxide solution. Initially a precipitate ($CaCO_3$) will form. Upon continued addition of carbon dioxide, the precipitate disappears due to formation of calcium bicarbonate, which is more soluble. This is a stoichiometry effect:

$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

$$CaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ca(HCO_3)_2(aq)$$

For further discussion see *Acids and Bases* module.

Analogies and Metaphors

Limiting Reactant

1. **Tricycle**: How many tricycles can be made from a given number of wheels, handle bars, frames, seats, *etc.*? What will be left? ***Example***: How many tricycles can you make from 10 wheels, 5 handle bars, 4 frames and 6 seats? *[3]*
2. **Making S'mores**: Purchase a package of six Hershey's™ chocolate bars, a box of graham crackers, and a bag of marshmallows.

   Recipe for making S'mores:

   3 chocolate squares
   1 graham cracker (broken into two parts)
   2 toasted marshmallows

   The package of chocolate bars contains 6 bars, each bar has twelve chocolate squares. Each box of graham crackers contains three packages of ten crackers (may vary). Each bag of large marshmallows contains 80 marshmallows (may vary).

   Ask students the following questions:

   a. What is the maximum number of S'mores that can be made? *[24]*

   b. Which ingredient is limiting? *[Chocolate]*

   c. How many each of the other ingredients will be left? *[6 graham crackers, and 32 marshmallows]*

   d. If we wanted to make more S'mores, how many more of the limiting ingredient would we need to buy? How many more S'mores could we make? *[We would need to buy two more chocolate bars. We could make six more S'mores.]*

   e. After the second purchase, what is the limiting ingredient, and what is left? *[The limiting ingredient is now the graham crackers. One half of a chocolate bar (6 squares), and 20 marshmallows would be left.]*

   If you feel brave enough and wealthy enough, you could bring the ingredients to class. Marshmallows can be toasted over a burner. (Marshmallow forks can be made from branches or wires. Low flames are recommended. If marshmallow drips inside the burner, a careful washing will remove it.)

   Safety Note: *Some teachers understandably feel that food should never be eaten in a laboratory setting. They point out that we can hardly emphasize this as a safety rule if broken occassionally.* This is an effective theoretical yield analogy without actually making and eating these treats in class.

3. **Making a BLT sandwich**: The number of sandwiches that can be made with each having three slices of tomato, four slices of bacon, one leaf of lettuce and two pieces of bread is limited by the number of slices of tomato, bacon and bread, and the number of lettuce leaves available. Students can determine the number of sandwiches (product) and the "leftovers" (remaining reactants) from a given quantity of each ingredient (reactants). (Other types of sandwiches could be used.)
4. **Clothes**: How many sets of clothes can be made from a given number of socks, shoes, slacks, shirts, *etc.*?
5. **Setting a table**: The number of place settings (of two spoons, one knife and two forks each) is limited by the number of knives, forks and spoons available. Pose problems such as : How many places can be set with 8 knives, 5 forks and 12 spoons? *[2]* What is the "limiting reactant"? *[Forks]* What flatware will be "leftover"? *[6 knives, 1 fork, 8 spoons]* This analogy was suggested by Robert C. Lamb.
6. **Vending machine**: If a vending machine will take several different denominations of coins, how many colas (or candy bars, chips, *etc.*) can be obtained from a given number of nickels, dimes and quarters? What coins remain? The cost of the item can be varied as well as the number of each type coin.
7. **Change machine**: If a given number of quarters is inserted in a change machine, how many dimes and nickels can be obtained, assuming 2 dimes and 1 nickel are given for each quarter inserted?

Pictures in the Mind

1. A balanced chemical equation can be visualized as a balanced see-saw or a balanced pair of scales (balance). Stoichiometry problems can be viewed as finding the balance point for the see-saw or scales by determining the masses of the remaining substances needed for balance when the mass of one substance is known.
2. Students should be reminded that you can't have more than 100% of anything. The percent yield cannot be over 100%! Visualize the theoretical yield as the maximum and the actual yield as something less than the maximum. Sometimes the percent yield is more than 100%. This could be due to impurities present contaminating the product or incorrect masses. Another possibility is that the proposed reaction is incorrect.

Other Activities

1. Batman and Robin in Chemistry: Stoichiometry problems based on comic strip (see *Appendix*).
2. Cork shooting contest: Students compete making the "best" cork shooting bottle, given certain parameters for quantities of sodium bicarbonate, acetic acid and bottle size (see *Appendix*).

TIPS FOR THE TEACHER

Language of Chemistry

stoichiometry relationships between quantities of substances (reactants and products) involved in a chemical reaction.

reactant substance that enters into chemical reaction, usually with one or more additional substances (also reactants). Reactants appear on the left in a chemical equation.

product substance formed as the result of a chemical reaction. Products appear on the right in a chemical equation.

Law of Conservation of Mass Mass (matter) is neither created nor destroyed in any chemical process. (In this module, the specific application is that mass is neither created nor destroyed in chemical reactions. The total mass of reactants is the same as the total mass of products in a chemical reaction. This law implies that the number of atoms, hence the moles, of each element is the same on both sides (reactants and products) of a chemical equation and that when an equation is balanced the number of each type atom is the same on each side of the equation.)

balanced equation represents a chemical reaction that conforms to the Law of Conservation of Mass. The molar amount of atoms of each element is the same on both sides (reactants and products) of a balanced chemical equation. Equations are balanced by placing a coefficient in front of a substance's formula to indicate the number of units of that substance required to bring the equation into balance.

mole ratio (also called stoichiometric ratio) fraction expressing the relationship of moles of one substance in a chemical reaction to moles of some other substance in the reaction. The mole ratio is derived from the coefficients in the chemical equation.

theoretical yield quantity of that product predicted from the stoichiometry of the reaction. (This statement assumes that the reaction goes to completion.)

actual yield quantity of product that is obtained when the reaction is carried out. The actual yield is never more than the theoretical yield and is frequently significantly less. The actual yield may be limited by the nature of the reaction, by the reaction conditions, by the inability to separate the desired product from remaining reactants or other products, or by the reactants producing "side products" by undergoing simultaneous competing reactions that do not yield the desired product.

percent yield fraction of the theoretical yield that is actually obtained, expressed as a percent.

$$\% \text{ Yield} = (\text{Actual yield}/\text{Theoretical yield}) \times 100$$

molarity (or **molar concentration**) moles of solute per liter of solution. This unit is used by chemists to express the concentration of solutions because the stoichiometry of chemical reactions is based on moles. If the concentration is expressed in molarity, the number of moles desired can be directly related to the volume (in liters) of the solution measured, facilitating stoichiometric calculations (see *Solutions* module).

Pattern Recognition

Coefficients in an equation relate moles of substances in a reaction. These coefficients are the numbers used in constructing the mole ratio for a stoichiometry problem.

Common Student Misconceptions

1. **"Subscripts can be changed to balance an equation."**

   Only coefficients can be changed when balancing equations. Changing subscripts changes the nature of the substances, if the "new formulas" represent real substances at all! The reaction is not the same even if mass balance is obtained.

2. **“The number of reactant moles must be the same as the product moles in a balanced chemical equation.”**

   This is generally not true because the atoms are arranged differently in the products and reactants. It is important to practice balancing equations where the number of moles of reactants and products differ.

3. **“The coefficients in the balanced equation relate the masses of the substances.”**

   It is important to emphasize that stoichiometric coefficients relate *moles* of substances in the balanced equation. The masses of the substances will vary because molar masses for different substances are not equal. (See section on *Problem Solving* for suggestions on how to deal with this misconception.)

4. **“Changing the coefficient for a substance changes the molar mass.”**

   Students may use the product of the molar mass and the coefficient as the molar mass. When the coefficient is also used in the mole ratio, the stoichiometry is incorrect.

5. **“Volumes of reactant solutions are related by the stoichiometric coefficients.”**

   This may be a problem if laboratory activites and demonstrations involve equal volumes of all reactants or volumes proportional to stoichiometric coefficients. Take care in selecting laboratory activities and demonstrations so that the solution volumes used are not always in the ratio of the moles of reactants.

6. **“The molar mass for hydrogen, oxygen, chlorine, *etc.* is the same as the atomic mass.”**

   In this module, the common diatomic gases should be reviewed. Students frequently use one-half the correct molar mass for hydrogen, oxygen, nitrogen and the halogens in stoichiometry problems because they assume all elements are monatomic.

Problem Solving

Balancing Equations

The following rules will provide guidance for students in balancing equations.

Five Commandments for Balancing Equations

1. Make a “score sheet” to keep track of elements in reactants and products. List elements in the order they appear in the reactants. Make an identical list for products. Record the number of atoms of each element for reactants and products. (If a polyatomic ion, such as NH_4^+ or NO_3^-, remains *unchanged* during the reaction, it can be treated in the same manner as an element.)

2. Fractional coefficients may be used. Coefficients can be converted to integers when the balancing is complete.

3. Use the following order for balancing the equation:

   a. Metallic elements
   b. Polyatomic ions that are *exactly the same* on both sides of the equation
   c. Nonmetallic elements that are *uncombined* on one side of the equation
   d. Remaining elements other than hydrogen or oxygen
   e. Hydrogen
   f. Oxygen

4. To remove fractional coefficients, multiply by an integer that will convert the fractions to integers. The integer should appear in the denominator of the fractional coefficient.
5. Using your original equation (*not the "score sheet"*), check that the equation is balanced (the same number of each atom type on both sides of the equation).

Solving Stoichiometry Problems

1. Emphasize the mole relationship in solving stoichiometry problems. Stoichiometry problems should be introduced by working mole-mole problems until the student is thoroughly familiar with the principle involved. The Mole A-Mole B conversion, and the mole ratio are the heart of stoichiometry problems.

   Then the mole-mass and mass-mole problem is introduced, again emphasizing that moles (not mass) of substances are related. Outlining the problem solution by making a "road map" is useful. For example, in the problem:

   "What mass of Al_2O_3 can be prepared by the reaction of 2.50 mol of Al according to the equation: $4\,Al + 3\,O_2 \rightarrow 2\,Al_2O_3$?"

   The "road map" would be:

   $$\text{Moles Al} \xrightarrow{\text{Mole ratio}} \text{Moles } Al_2O_3 \xrightarrow{\text{Molar mass } Al_2O_3} \text{Mass } Al_2O_3$$

   $$\text{Mole ratio} = \frac{2 \text{ mol } Al_2O_3}{4 \text{ mol Al}}$$

   The "road map" states steps involved in solving the problem. Start with the known (given) quantity and devise a plan to obtain the desired quantity. For each step, state the information needed to make the "conversion." Emphasize the centrality of the Mole A to Mole B conversion when working mass-mole and mole-mass problems.

   If sufficient practice is provided with these two types of stoichiometry problems, the transition to mass-mass problems (the most common and practical type problem) should be smoother, and students should have fewer problems. This type of problem adds an additional step to the "road map," but the step is the same as has already been encountered. There is a mass-mole conversion on "both ends" of the mole-mole conversion. Be sure to tell the students when to stop (when the desired unit is obtained).

2. Most texts use the factor-label method (dimensional analysis) for solving stoichiometry problems. Emphasize the importance of:
   a. Keeping the formula for the substance and units with each number.
   b. Outlining the "road map" as a guide to setting up the problem.
   c. All stoichiometry problems involve a mole ratio (to convert moles of one substance into moles of another substance).
   d. The molar mass relates mass to moles.
   e. The molarity relates moles to volume, in liters, of a solution.
   f. The mole ratio involves "counting numbers," and thus, does not limit the significant figures in the answer.
   g. The molar mass should be determined to at least as many significant figures as the given (or known) quantity in the problem.

   h. After setting up the problem, check that units cancel to give the desired unit before manipulating the numbers in the problem. If the answer has incorrect units, the numerical part will not be correct either.

3. Most beginning students find it easier to understand limiting reactant problems if they calculate the yield of product from the given quantity of *each* reactant. They then compare the yield from each reactant. The maximum yield for the reaction (theoretical yield) will be the *smallest* of the calculated yields (see *Metaphors and Analogies*).

   Since the mass of product is usually desired, one of your answers (the theoretical yield) will be the answer to the problem. If the quantity of "leftover" reactants is desired, calculate the quantity of each reactant that will combine with the limiting reactant. Subtract the amount that reacts from the given amount to determine the quantity of "leftover" reactant.

4. It is important to generalize the approach to the different types of stoichiometry problems so that students begin to see the pattern in the solution:
   a. Given data about A, find Mole B.
   b. Relate Mole A to Mole B.
   c. From Mole B, find desired information about B (if answer has units other than Mole B).

   The mole ratio is the heart of stoichiometry. The previous approach will be used when gas reactions are studied and in the discussion of ionic equilibrium, electrochemistry, *etc.* It is central to all topics involving chemical reactions.

   Although stoichiometry problems involving heat may be discussed at this point or in the *Thermochemistry* module, reviewing stoichiometry at a later point in the course emphasizes the relationship of this key chemical concept to other topics.

Decision Making

It is important to stress the concept that "you can't get something from nothing" and "matter never disappears." This applies to chemistry (Law of Conservation of Mass) as well as to life itself. Pollution and waste disposal problems facing communities and nations are linked to stoichiometry. Students should realize that chemical reactions can be used to change the form of matter (*e.g.*, change a harmful or useless substance into something less harmful or even something useful), but it is impossible to make pollution or waste "disappear." Factors to consider, in addition to the feasibility of the process, are cost (economics), environmental effects of the clean-up (sometimes the by-products of clean-up are harmful), public perception of the problem, and the proposed solution.

Ask students to assume that they are managers in a chemical manufacturing plant. Discuss the factors to be considered in their responsibilities to their employer and society. Factors include alternative manufacturing processes (different chemical reactions with different starting materials to produce the same product—economics and stoichiometry are important issues here), source and consumption of raw materials (calculation of amounts of materials using stoichiometry is important here as well as economic and environmental considerations), waste disposal, pollution and public health issues (again economics and stoichiometry enter the picture). Students should be guided to see that there is no simple "right answer" here. Like other decisions in our lives, there are tradeoffs. As responsible citizens, we must make informed decisions concerning complex issues. The study of chemistry prepares us to make those decisions.

HISTORY: ON THE HUMAN SIDE

Scientists who contributed to our understanding of the quantitative relationships in chemical reactions include:

1. **Joseph Louis Gay-Lussac** (1788-1850, France). His experiments led him to propose in 1808 the Law of Combining Volumes, which states that the volumes of gases involved in a chemical reaction are in a small whole number ratio.
2. **Amedeo Avogadro** (1776-1856, Italy). In 1811, the Avogadro Hypothesis (equal volumes of gases at the same temperature and pressure contain equal numbers of molecules) was deduced from his observations of gas reactions and those reported by Gay-Lussac.
3. In the period 1949-1953, **E. Chargaff** and co-workers concluded that the mole ratio in DNA for adenine: thymine and cytosine: guanine was always 1:1, thus paving the way for the discovery of the precise structure of DNA by J. Watson and F. Crick.

HUMOR: ON THE FUN SIDE

1. Word Search (see *Appendix* for master copy)

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A N G D P I U B Z U A C H E K E R
S T P I R C S B U S V Y P D L H Q
T P Y V M B O E I F D C L O I Y U
N G V Z L O Q E B T R D M L M T P
E W Q X N S L Z N I P I G A I V N
I V P P D V B A L A N C E D T V L
C E L A Y K T S R Q X U F R I R X
I S K J I C C C T I X S N I N J W
F W G K A E W P O R T W M H G N Z
F N X E Y W X D K Y H Y O O H V P
E P R O D U C T C F N U C I T U H
O I O J B E N H D D L A U T C A T
C B J W F D P P D Y B B K L I D V
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Substance that normally appears on the left side of a chemical equation.
2. Substance that appears on the right side of a chemical reaction.
3. These are always conserved in a chemical reaction.
4. Adjective describing the yield of product obtained when a reaction is carried out.
5. Number of moles of solute per liter of solution.
6. Chemical equations are balanced with these numbers.
7. Numbers never changed when balancing a chemical equation.
8. Type of reactant present in smallest stoichiometric amount is said to be this.
9. If a chemical expression conforms to the Law of Conservation of Mass, it is said to be ____.
10. The _________ ratio is the heart of stoichiometry.

Answers: 1. REACTANT 2. PRODUCT 3. ATOMS 4. ACTUAL 5. MOLARITY 6. COEFFICIENTS 7. SUBSCRIPTS 8. LIMITING 9. BALANCED 10. MOLE

2. See cartoons at end of module.

MEDIA

1. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
2. *Stoichiometry,* by C. A. Wilkie. Saunders Computer Chemistry Apple IIe Software Package.
3. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, Wl 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
    a. For the Apple II computer running on ProDOS: AR 301, AR 302.
    b. For the Apple II computer: AP 301, AP 304, AP 305, AP 307.
    c. For IBM PCs and PC-compatibles: PC 2201, PC 2301, PC 2302.
4. “The Atom” from the Ring of Truth series by Philip Morrison includes about a 3-minute segment on the reaction of different volume ratios of gases. This series appeared on PBS television during 1987. Videotapes are available from PBS Home Video (Phone 800-424-7963).
5. “Changes” from the Ring of Truth series by Philip Morrison deals with chemical reactions as well as heat change. This series appeared on PBS television during 1987. Videotapes are available from PBS Home Video (800-424-7963).
6. TV Ontario has a videotape series that contains short videos that could be used in this module. Free previews are available. TV Ontario Video, 143 West Franklin Street, Suite 206, Chapel Hill, NC 27516.
7. Doing Chemistry Videodiscs are available from the Education Division, American Chemical Society, 1155-16th Street, NW, Washington, DC 20036.
8. Software published by *JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, Wl 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
    a. *ChemCalc: A Scientific Calculator,* by Robert Allendoerfer. Vol. III B, No. 2, for IBM PS/2 PC-compatible computers.
    b. *REDOX: A Tutorial on REDuction / OXidation Equations,* by Derek Davenport, Paul Groves and Dale Jensen. Vol. III A, No. 1, for the Apple II computer.
    c. *Bomb Calorimeter Simulation,* by David J. Olney. Vol. III A, No. 2, for the Apple II computer.
9. Videodisc published by *JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, Wl 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

    “The Mole–From Counting to Weighing” and “Blowing Balloons with Moles of Hydrogen,” two chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.

INSTRUMENTATION

The *Electrochemistry* module has information on electrolysis and construction of electrolysis apparatus.

Links/Connections

WITHIN CHEMISTRY

1. **Acids and Bases**. The reaction of acids and bases to produce water and a salt is called a neutralization reaction. The number of hydrogen ions (from the acid) neutralized in the reaction is equal to the number of hydroxide ions (from the base) that react. Complete neutralization occurs when all H^+ ions from the acid and all OH^- ions from the base react. Many neutralization reactions are carried out in solution. If the acid and base are both strong electrolytes, the net ionic equation for the reaction is:

   $$H^+(aq) + OH^-(aq) \rightarrow HOH(l)$$

   Titration is the procedure used to determine the concentration of a solution by reaction with another solution of known concentration or with a known mass of a solid. Titrations are frequently employed not only in the chemistry laboratory, but in industrial analytical laboratories (see *Acids and Bases* module).
2. **Chemical Equilibrium**. Writing a chemical equation often implies that reactants are completely consumed in forming products, but this is not always the case. Often there are significant amounts of both reactants and products when the reaction has apparently ceased (no measurable change in amounts of reactants and products can be detected). In such cases, the reaction has reached chemical equilibrium. The equilibrium concentrations of reactants and products can be used to determine the equilibrium constant, a quantity relating concentration of reactants and products present at equilibrium. For a reaction with a known equilibrium constant, the concentrations of all species present at equilibrium can be determined by knowing the amounts of substances added to the reaction vessel. Stoichiometry plays an important role in determining the concentrations of species present at equilibrium, since the coefficients in the balanced chemical equation are used in the mathematical relationship used to evaluate equilibrium concentrations (see *Equilibrium* module).
3. **Electrochemistry**. The amount of chemical change is related to the number of electrons transferred in an electrochemical reaction. The amount of chemical change in an electrolysis reaction is also related to current and time. For electrochemical reactions, stoichiometry problems may involve moles of electrons as well as reactants and products (see *Electrochemistry* module).
4. **Gas Stoichiometry**. Since one mole of an ideal gas occupies the same volume as a mole of another ideal gas at the same temperature and pressure, the volumes of reacting and product gases in a chemical reaction are in the same proportion as the stoichiometric coefficients. This relationship is referred to as Gay-Lussac's Law of Combining Volumes. (It should be noted that this relationship only holds for *gases*, not liquids or solids.) This relationship facilitates the determination of reactant and product volumes for gases (see *Gases* module).
5. **Net Ionic Equations**. Net ionic equations are important in solution chemistry because they show the "essential chemistry" of a reaction. In a net ionic equation, all ions not directly involved in the reaction ("spectator ions") are omitted. Frequently net ionic equations are more easily balanced than molecular equations. This is particularly true of oxidation-reduction reactions.

6. **Molarity and Solution Stoichiometry**. Many chemical reactions take place in solution. In these cases, concentration, rather than solute mass, is generally known for reactants. Since stoichiometric coefficients relate moles of reactants, the most useful concentration unit is molarity (moles solute per liter solution). Because net ionic equations are frequently used to describe reactions taking place in solution, it is important to be able to associate reactants, for which compound formulas are generally given, with the constituent ions, which appear in the net ionic equation. The stoichiometric coefficients obtained from the net ionic equation will be used in conjunction with molarity and/or molar mass in predicting quantities of reactants needed or products formed (see *Solutions* module).
7. **Thermochemistry**. Chemical reactions involve absorption or evolution of heat. The quantity of heat involved in a chemical reaction is directly related to the amounts of reactants and products (see *Thermochemistry* module).

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

1. **Chemical Engineering**. Commercial chemical manufacturing processes (such as the Contact process for sulfuric acid, the Ostwald process for nitric acid, and the Haber process for ammonia) involve chemical reactions. In selecting a reaction for industrial manufacture, consideration is given to yield as well as cost and availability of raw materials and adaptability to mass production (see *Industrial Inorganic Chemistry* module).
2. **Environmental Science**. An understanding of the environment requires knowledge of stoichiometry. For example, aerobic bacteria convert organic pollutants in streams to carbon dioxide and water. As the process occurs, dissolved oxygen is consumed, depleting the water of the oxygen required for the survival of marine life. Determination of the effects of dumping of organic chemicals (including sewage) into streams involves stoichiometry of the reactions involved.
3. **Geology**. Formation of stalactites and stalagmites involves the dissolution and reprecipitation of $CaCO_3$. Dissolved CO_2 causes limestone ($CaCO_3$) to dissolve according to the reaction:

   $$CaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ca^{2+}(aq) + 2HCO_3^{-}(aq)$$

   The reverse reaction results in reprecipitation of $CaCO_3$. (The reaction of $CaCO_3$ and $MgCO_3$ with dissolved CO_2 is also responsible for producing hardness in natural water supplies. See *Rocks, Minerals and Gems* and *Acids and Bases* modules.)
4. **Materials Science**. Relative proportions of starting materials and reaction conditions affect properties of polymeric materials. Most polymeric materials do not have fixed composition, that is, they contain molecules of varying molar mass but approximately the same proportions of constituent monomers. Properties (such as strength, resistance to chemical attack) of polymeric materials are correlated to reaction conditions to develop improved polymers. Glasses, ceramics, and metal alloys also are prepared with varying proportions of additives to alter the properties (see *Polymers* and *Materials Science* modules).
5. **Metallurgy**. The potential yield of metals from their ores can be determined from the reaction stoichiometry used in metallurgical processes. The ore's composition, as well as production costs and market demand determine the feasibility of development of mining operations of newly discovered mineral deposits.

6. **Biochemistry and Medicine**. Clinical laboratory scientists use stoichiometry in the study and analysis of body fluids. Metabolism involves the conversion of one substance into other substances. For example, in a known metabolic reaction, a product's (called a metabolite) concentration of a reaction in blood or urine can be used to determine the concentration of the original substance (see *Organic Chemistry, Chemistry in Medicine, Forensic Chemistry,* and *Enzymes: Biochemical Catalysts* modules).

TO THE CONTEMPORARY WORLD

Personal

Careers as research chemist, laboratory technician, medical technologist (clinical laboratory scientist), chemical manufacturer (including quality control as well as plant operations) involve the use of stoichiometry.

Community

1. **Field trips**: water treatment/sewage disposal plants, chemical industry, environmental testing laboratory, medical laboratory (clinical laboratory).
2. **Knowledgeable individuals**: chemical engineers, clinical laboratory scientists (medical technologists), environmental scientists/hygienists, industrial chemists, pharmacists, recycling engineers (resource management), research chemists, sanitation department chemists/water chemists, textile chemists

Societal (Science/Technology/Society; Current Events)

Stoichiometry is related to many important issues in everyday life and topics of current interest. A few of these are listed below.

1. **Comparison of Heat from Combustion of Various Fuels**. Fuels, such as coal, natural gas and petroleum, produce different amounts of heat energy upon combustion. Fuels, such as alcohols, that can be produced from renewable resources like grain or trees have been investigated as future energy sources. Electrical energy, solar energy, geothermal energy, and nuclear energy are also alternative energy sources. Comparison of the practicality of various energy sources for specific purposes involves many factors. One of these is combustion energy. Stoichiometry and thermochemistry provide the basis for the calculation of energy available from the fuel. Other factors are also important, however. Natural gas (methane) is a practical fuel for fleet vehicles (such as buses) that can return to a common refueling site each day but is not practical for private vehicles because of the complexity of refueling (see *Thermochemistry* module).
2. **Air Pollution and Fossil Fuel Combustion**. Burning coal or other fossil fuels containing sulfur or sulfur compounds generates sulfur dioxide, which produces acid rain (H_2SO_4) as shown in the series of equations:

$$S(s) + O_2(g) \rightarrow SO_2(g)$$
$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$
$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

Because of high temperatures reached in the internal combustion engine, nitrogen and oxygen from the air can react to form nitrogen oxides, also associated with air pollution. In addition to other effects on the atmosphere, nitrogen oxides also contribute to acid rain (HNO_3) as shown in the series of equations:

$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$
$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$
$$2NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + HNO_2(aq)$$

Incomplete combustion of organic fuels (C_8H_{18}) can result in the formation of carbon monoxide, another air pollutant.

$$2C_8H_{18}(l) + 17O_2(g) \rightarrow 16CO(g) + 18H_2O(l)$$

3. **Ozone Depletion**. Freons (chlorofluorocarbons) released into the atmosphere from refrigerators, freezers, and air conditioners, result in photochemical reactions in the stratosphere. One reaction involves ozone destruction. Since ozone is responsible for absorbing ultraviolet light before it reaches the earth's surface where it increases the incidence of skin cancer in humans, ozone depletion has serious health implications (see*Photochemistry* module).
4. **Water Purification and Waste Water Treatment**. The first step in the purification of city water supplies involves allowing water to settle in large reservoirs. To facilitate this process, lime, and aluminum sulfate are added to produce gelatinous aluminum hydroxide, which settles slowly, carrying suspended materials, including bacteria, with it. Remaining bacteria are killed by addition of chlorine to the water.

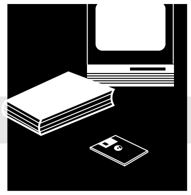
$$3Ca(OH)_2 + Al_2(SO_4)_3 \rightarrow 2Al(OH)_3 + 3CaSO_4$$

This same process is used in the primary treatment of wastewater. Secondary and tertiary treatment of wastewater involves bacterial oxidation to convert soluble organic materials to substances that will settle, absorption of chemicals by activated carbon, and steam distillation of volatile organic pollutants.

Extensions

Batman and Robin in Chemistry (see *Appendix*)
Cork Shooting Contest (see *Appendix*)
Activity 1: Limiting Reagent for a Chemical Reaction in Solution in *Solutions* module

References



Module developed by Caroline L. Ayers, Philip Dail, and Carolyn Morse, the North Carolina team.

Acid-Base Reaction Saves Batman's Life. (1974). *Chemistry, 47*(9), 5.

The activity involving the Batman cartoon is based on this article and an adaptation supplied by Larry Duberich.

Cotton, F. A., Darlington, C. L., and Lynch,C. L. (1980). *Chemistry: An investigative approach*. Boston, MA: Houghton Mifflin.

This text contains many good student activities on stoichiometry.

Dail, P. (1984). *High school chemistry: A laboratory approach*. Raleigh, NC: Wake County Board of Education.

The percentage yield of a chemical reaction activity is adapted from a similar experiment given in this laboratory manual.

Doing Chemistry [Videodisc Program]. Washington, DC: Education Division, American Chemical Society.

Segments illustrate demonstrations, experiments, and techniques. Consult the *Teacher's Manual* supplied with the videodisc to locate segments.

Fenster, A. E., Harpp, D. N., Schwarcz, J. A., and Glanville, J. O. (1988). A well-known chemical demonstration to illustrate an unusual medical mystery. *Journal of Chemical Education, 65*, 621.

This article briefly describes a medical condition caused by failure of the body to reduce Fe^{3+} to Fe^{2+} in the blood and the treatment for the condition. An analogy is provided by the "blue bottle" demonstration, which also uses methylene blue in an oxidation-reduction reaction.

Gardner, M., Smith, P., and Heikkinen, H. (1990). *Laboratory assessment builds success*. Berkeley, CA: Lawrence Hall of Science, Institute for Chemical Education.

This laboratory manual contains tested student activities written in the same format as those in SourceBook.

Morrison, P., and Morrison, P. (1987). *The ring of truth: An inquiry into how we know what we know*. New York, NY: Random House.

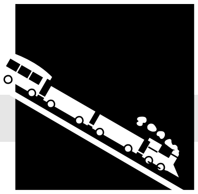
This book contains discussions about atoms, reactions, *etc.* that might be used in conjunction with this module.

Ramette, R. W. (1980). Exocharmic reactions. *Journal of Chemical Education, 57*, 68-69.

This article discusses experiments and demonstrations as motivational devices.

Szabadvary, F. (1972). Great moments in chemistry. Part VII. Chemistry learns to calculate (R. E. Oesper, Trans.). *Chemistry, 45*(4), 10-12.

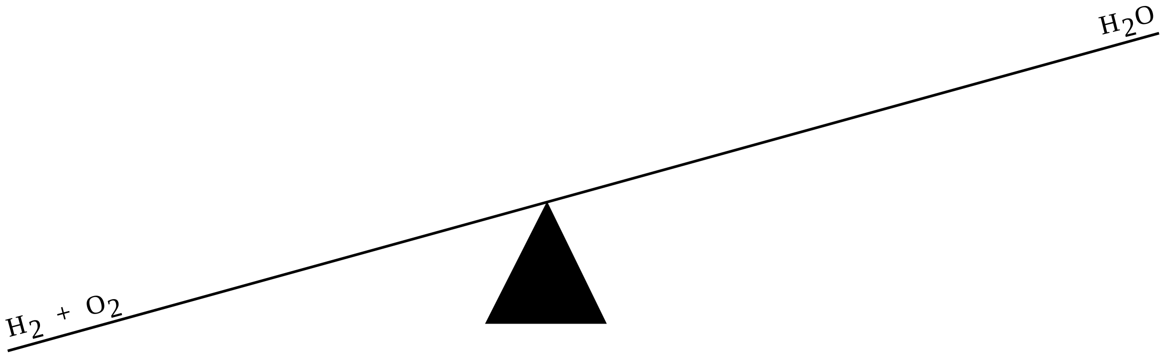
This is a play from a 15-part series originally presented over Hungarian radio and television. A reporter and a professor make a time reversal visit to Jeremias Benjamin Richter (1795) to discuss his early contributions to stoichiometry.

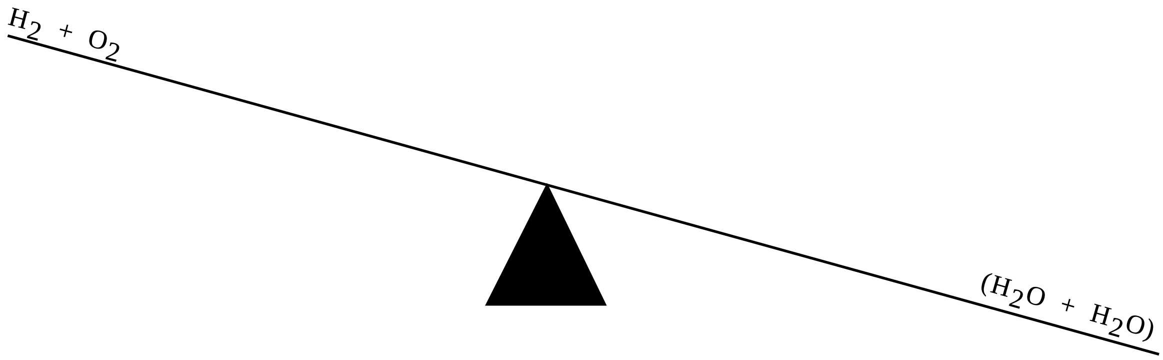


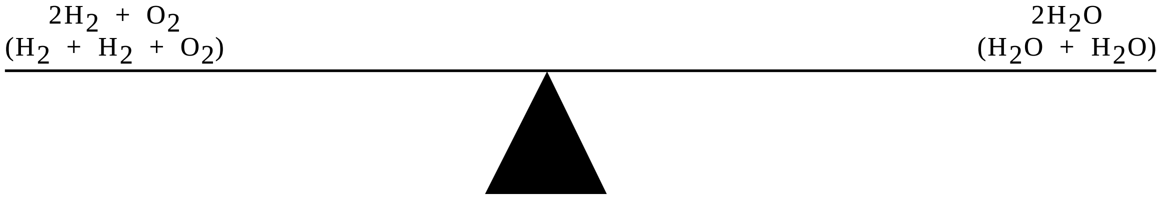
Appendix

- **Transparency Masters**
  1. Equation Balance
  2. Word Search
- **Humor**
  1. Batman and Robin in Chemistry
  2. Drawing of cylinder tank showing dimensions in Batman and Robin Chemistry
  3. Solutions to Batman and Robin in Chemistry
  4. Additional Activity: Cork Shooting Contest
  5. Cartoons

“Equation Balance”







Word Search

A	N	G	D	P	I	U	B	Z	U	A	C	H	E	K	E	R
S	T	P	I	R	C	S	B	U	S	V	Y	P	D	L	H	Q
T	P	Y	V	M	B	O	E	I	F	D	C	L	O	I	Y	U
N	G	V	Z	L	O	Q	E	B	T	R	D	M	L	M	T	P
E	W	Q	X	N	S	L	Z	N	I	P	I	G	A	I	V	N
I	V	P	P	D	V	B	A	L	A	N	C	E	D	T	V	L
C	E	L	A	Y	K	T	S	R	Q	X	U	F	R	I	R	X
I	S	K	J	I	C	C	C	T	I	X	S	N	I	N	J	W
F	W	G	K	A	E	W	P	O	R	T	W	M	H	G	N	Z
F	N	X	E	Y	W	X	D	K	Y	H	Y	O	O	H	V	P
E	P	R	O	D	U	C	T	C	F	N	U	C	I	T	U	H
O	I	O	J	B	E	N	H	D	D	L	A	U	T	C	A	T
C	B	J	W	F	D	P	P	D	Y	B	B	K	L	I	D	V

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Substance that normally appears on the left side of a chemical equation.
2. Substance that appears on the right side of a chemical reaction.
3. These are always conserved in a chemical reaction.
4. Adjective describing the yield of product obtained when a reaction is carried out.
5. Number of moles of solute per liter of solution.
6. Chemical equations are balanced with these numbers.
7. Numbers never changed when balancing a chemical equation.
8. Type of reactant present in smallest stoichiometric amount is said to be this.
9. If a chemical expression conforms to the Law of Conservation of Mass, it is said to be ____.
10. The _________ ratio is the heart of stoichiometry.

Batman and Robin in Chemistry

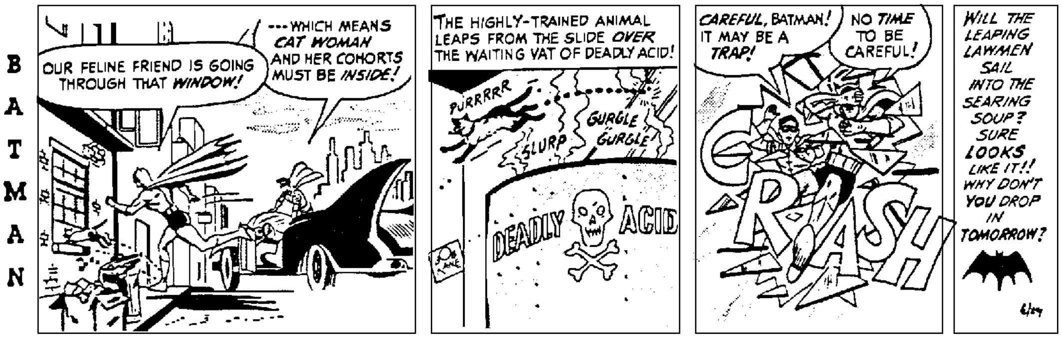
(Adapted from activity supplied by Larry Duberich. Used by permission of American Chemical Society. See first reference in *References* section.)

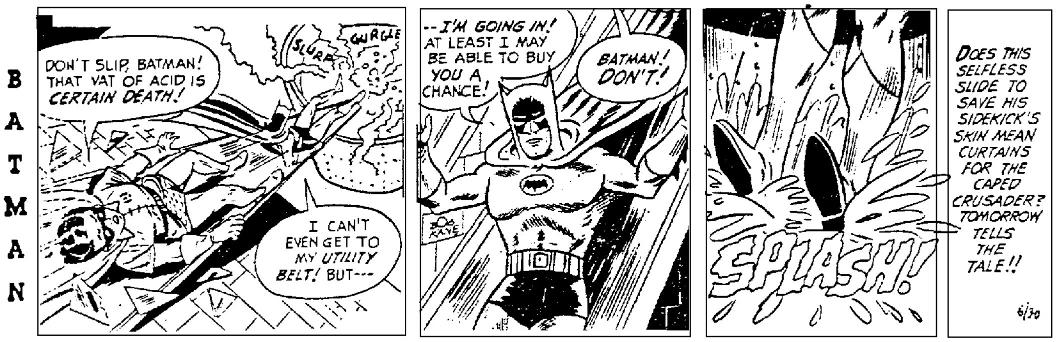
Batman uses pellets, which he does not identify, to neutralize the acid in the tank into which he falls. He does not give the size of the tank. Suppose the pellets were NaOH and the tank was 3.00 meters in diameter filled with 98% sulfuric acid by mass to a depth of 1.25 meters (see transparency).

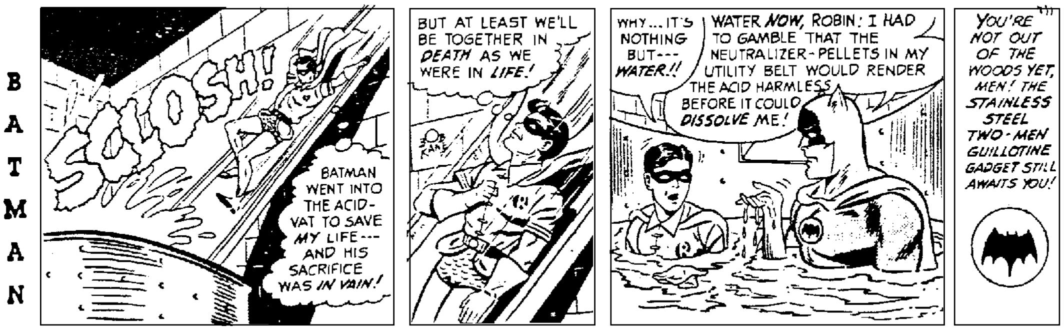
Perform the following calculations and *show your work*.

1. Calculate the volume of acid (in cm^3) in the tank. (The volume of a cylinder is $\pi r^2 h$.)
2. Calculate the mass of acid in the tank, then find the number of moles of H_2SO_4 present. (Concentrated sulfuric acid is 98% (by mass) H_2SO_4 and has a density of 1.84 g/cm^3.)
3. One mole H_2SO_4 produces *two* moles of H^+ ions. With this in mind, write the balanced equation for the neutralization of NaOH and H_2SO_4 . Determine the number of moles NaOH required to neutralize the acid in the tank. What *mass* of NaOH is required?
4. The heat of reaction for neutralization of a strong acid and a strong base is –57.3 kJ for each mole of H_2O formed. How much thermal energy (in kJ and J) was released in the neutralization reaction?
5. If the initial temperature of the acid bath was 20 °C, how much thermal energy would be needed to raise the temperature of the mixture in the tank (NaOH + H_2SO_4) to 100 °C? Assume the mass of the solution is the combined masses of the H_2SO_4 and NaOH. Also, assume the heat capacity of the liquid in the tank is 4.184 J/g.
6. Was all the generated thermal energy (Question 4) used to heat the liquid in the vat to 100°C (Question 5)? After water reaches the boiling point, an additional 2260 J/g is required to vaporize it. How much water was evaporated by the heat generated in the reaction of NaOH with H_2SO_4?

   (Ignore the boiling point elevation.)
7. In view of answers to Questions 3 and 6, how likely is it that the comic strip author passed high school chemistry?







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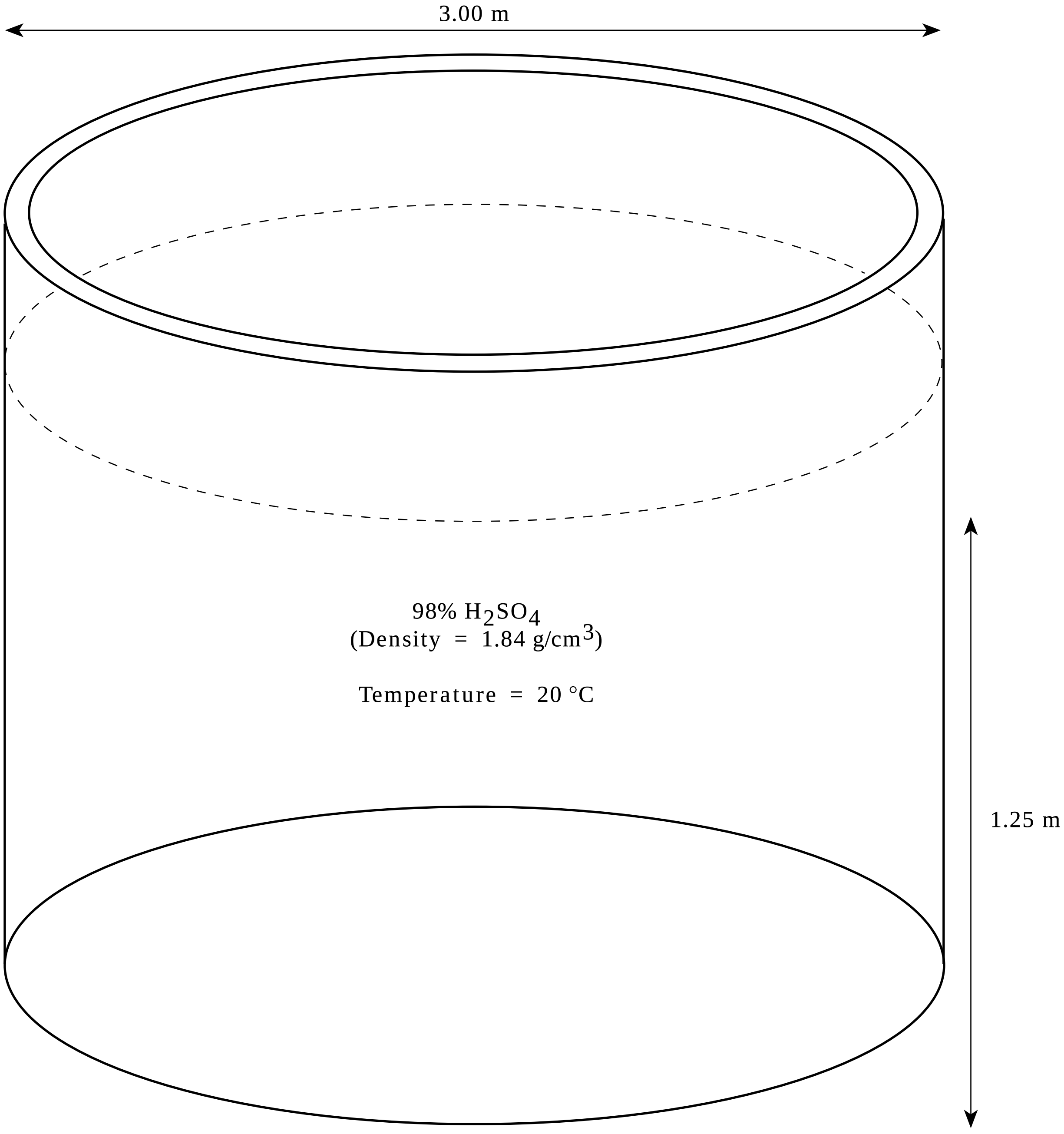
Cylindrical Tank



Solutions to Batman and Robin in Chemistry

1. Tank diameter = 3.00 m, r = 1.50 m, $r^2 = 2.25\ m^2$, h = 1.25 m
   $V = 8.84\ m^3 = 8.84 \times 10^6\ cm^3\ [1\ m^3 = (10^2\ cm^3)^3]$

2. $8.84 \times 10^6\ cm^3 \times 1.84 \frac{g}{cm^3} = 1.63 \times 10^7 g$ acid

   $1.63 \times 10^7 g\ acid \times \frac{98}{100} \times \frac{1\ mol}{98.1\ g} = 1.63 \times 10^5 mol\ H_2SO_4$

3. $H_2SO_4 + 2NaOH \rightarrow 2H_2O + 2Na^+ + SO_4^{2-}$

   $1.63 \times 10^5\ mol\ H_2SO_4 \times \frac{2\ mol\ NaOH}{1\ mol\ H_2SO_4} \times \frac{40.0\ g\ NaOH}{1\ mol\ NaOH} = 1.30 \times 10^7\ g\ NaOH$

   $1.30 \times 10^7\ g = 1.30 \times 10^4\ kg = 1.30 \times 10^1$ or 13.0 metric tonnes of NaOH

4. $1.63 \times 10^5\ mol\ H_2SO_4 \times \frac{2\ mol\ H_2O}{1\ mol\ H_2SO_4} \times \frac{57.3\ kJ}{1\ mol\ H_2O} = 1.87 \times 10^7\ kJ$

   $= 1.87 \times 10^{10}$ J released

5. Assume: solution mass = mass acid + mass NaOH

   1.63×10^7 g acid + 1.30×10^7 g NaOH = 2.93×10^7 g solution

   if initial temp = 20 °C, final temp = 100 °C

   heat absorbed = $2.93 \times 10^7\ g \times 80°C \times 4.184 \frac{J}{g\ °C} = 9.81 \times 10^9\ J$

6. Reaction produces: 18.7×10^9 J (1.87×10^{10} J)

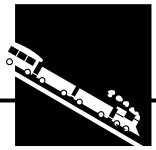
   Heating to 100 °C uses 9.8×10^9 J. 18.7×10^9 J released – 9.8×10^9 J used = 8.9×10^9 J remain.

   $8.9 \times 10^9\ J \times \frac{1\ g}{2260\ J} = 3.9 \times 10^6$ g solution are vaporized

   $\frac{3.9 \times 10^6\ g}{2.93 \times 10^7\ g}$ = 13% of the solution in the tank is vaporized

   (Note: The boiling point of a solution is greater than that of the solvent. This is ignored in these calculations.)

7. 13 metric tonnes of NaOH sounds like a serious Bat–Hernia. Furthermore, the Dynamic Duo were certainly cooked!



Cork Shooting Contest Rules and Regulations

(Adapted from activity submitted by John Garner)

Purpose

To determine what method of mixing sodium bicarbonate and acetic acid will shoot the cork the farthest.

Reaction

$NaHCO_3(s) + HC_2H_3O_2(aq) \rightarrow CO_2(g) + H_2O(l) + NaC_2H_3O_2(aq)$

Materials

- Cork
- Plastic Pop Bottles
- Protractor
- Tape measure
- Baking Soda (Sodium bicarbonate, $NaHCO_3$)
- Vinegar (5% acetic acid, $HC_2H_3O_2$)

Procedure

This is a student project involving concepts of chemistry and physics. *Students should observe standard safety procedures.* Design the procedure such that the optimum amounts of baking soda and vinegar are used to shoot the cork the maximum distance. Additionally, consider the bottle size and angle of tilt to maximize the cork's flight distance.

Safety

Use a plastic bottle. A glass bottle might explode during the reaction creating a safety hazard. (If a glass bottle must be used, it should be wrapped in heavy tape to prevent shattering in case of explosion.)

Rules

1. Must be done outside.
2. No more than two tablespoons baking soda ($NaHCO_3$).
3. No more than two tablespoons of vinegar ($HC_2H_3O_2$).
4. Cork must be decorated and named.
5. Data from 10 trials must be turned in before you can enter the contest.
6. Use 6, 10, 12, 16 or 32 ounce bottles only.
7. Bottle must be held on the ground.

Variables

1. Size of bottle
2. Quantity of vinegar
3. Angle of bottle
4. Quantity of baking soda
5. Design of cork
6. Mass of cork
7. Extent of shaking of bottle
8. Position of cork in the bottle

Information to Include in the Write-up

1. What makes the cork fly out of the bottle?
2. How many grams are in a tablespoon of vinegar? How many grams are in a tablespoon of baking soda?
3. Complete the tables of your trials at home and the trials of all your classmates from school during the competition at school.
4. What problems did you encounter and how did you solve them?
5. From all the data predict what the best variable combination is to get the longest distance.
6. Make a graph of distance *vs.* bottle size.
7. Make a graph of distance *vs.* amount of vinegar.
8. Make a graph of distance *vs.* amount of baking soda.

Report

1. Measurements: Distance cork travels
2. Data and Calculations: Questions 3, 6, 7, 8
3. Answer: Questions 1, 2, 4, 5

Date:

Name:

Home Laboratory - Cork Shooting Contest

Personal Data Sheet

To earn points, this worksheet may be turned in by:

____________________+ 5 ____________________ + 3 ____________________ + 1

These trials must be completed at home and turned in before: _______________

Each student should run at least 10 trials at home and record all the data below.

Name:

Design of Cork A

Name:

Design of Cork B

Trial	Cork Name	Volume Acid	Volume Soda	Type Bottle	Size Bottle	Angle of Bottle	Distance Traveled
1							
2							
3							
4							
5							
6							
7							
8							
9							
10							

This information should be included in the data and calculation section of your formal laboratory write-up.

Date:

Class Data Sheet

Weather:

Trial	Cork Name	Volume Acid	Volume Soda	Type Bottle	Size Bottle	Angle of Bottle	Distance Traveled
1							
2							
3							
4							
5							
6							
7							
8							
9							
10							
11							
12							
13							
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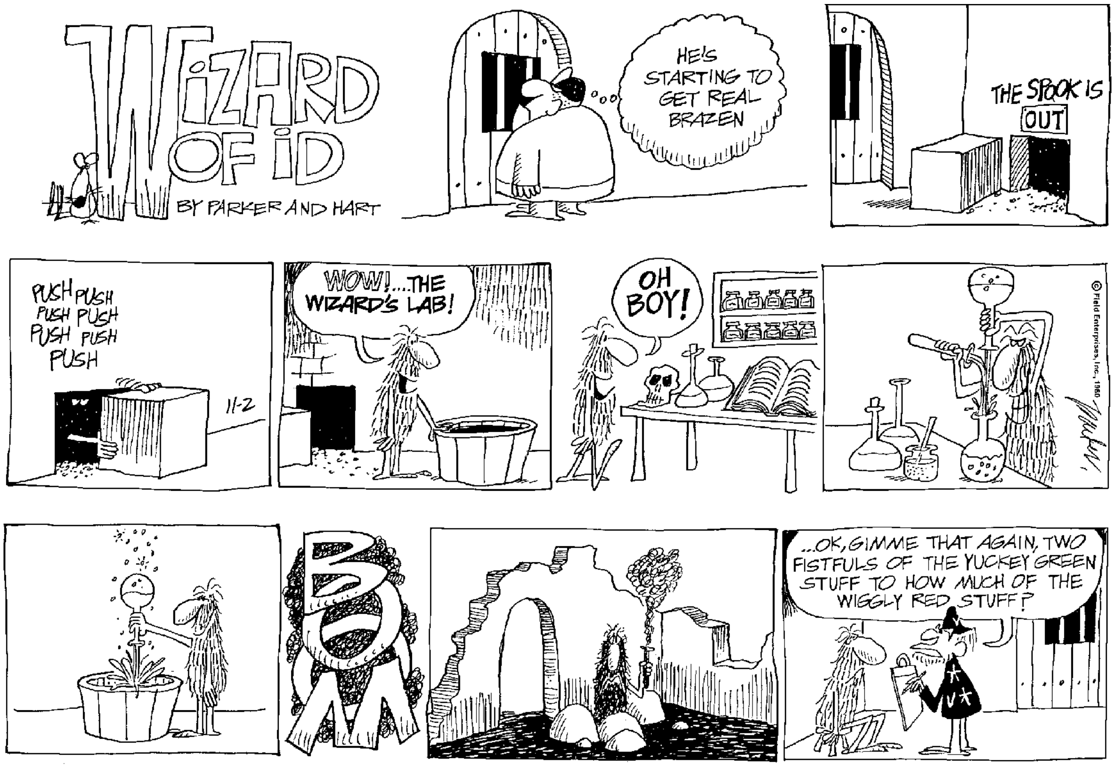
18							
19							
20							



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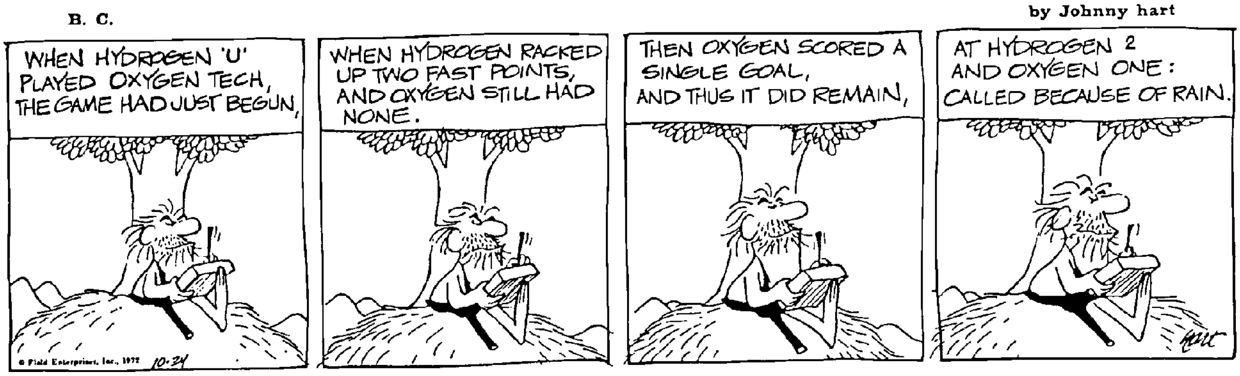

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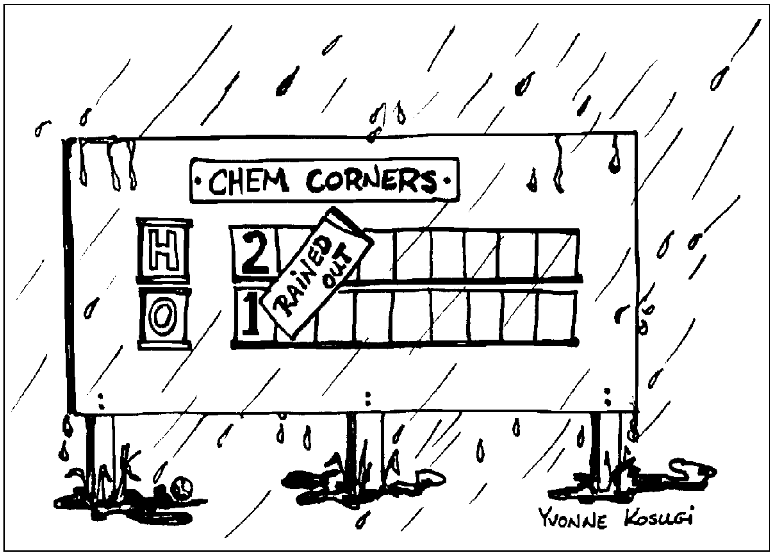

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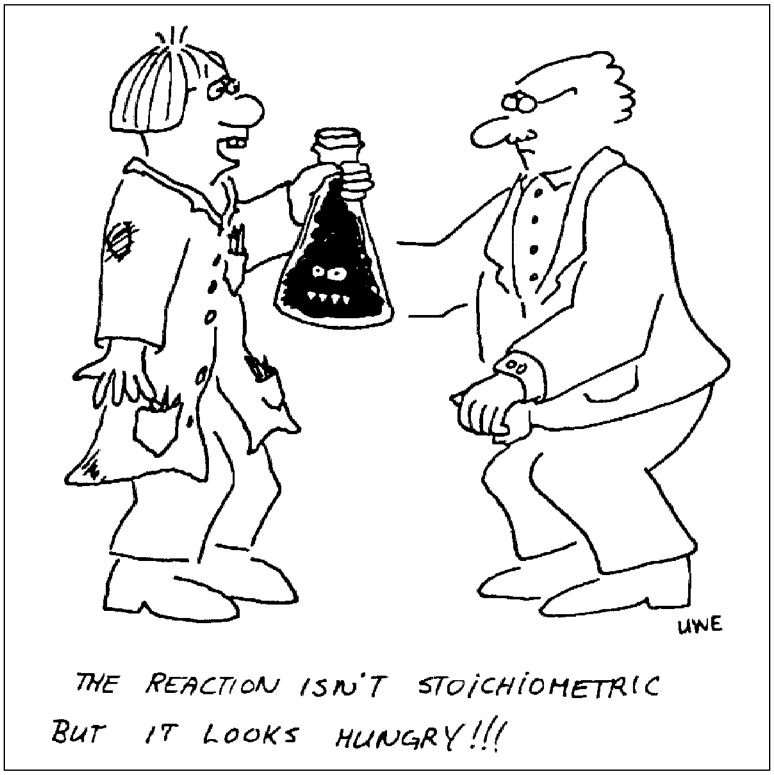

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