Chemistry 20

Unit D Quantitative Relationships in Chemical Changes

Module 6—Stoichiometry

Module Introduction

During your study in this course you have seen that chemical reactions are important to the function of the human body, to interactions in the biosphere, and to many technologies used today.

In this module you will study chemical reactions from a different perspective by looking at the amounts of each substance involved in the reaction. Investigating quantitative relationships involves asking questions like

- How much will react?
- How much will be produced by the reaction?
- How complete was the reaction?

Although understanding the amount of matter involved in a chemical process is one of the more obvious aspects to consider when studying chemistry, you can only begin to ask it now. This unit will test your ability to manage and manipulate numerical data. You will also use numbers to help you keep track of the amounts of matter you are investigating.

- Think about the following questions as you complete this module:
- How do scientists and engineers use mathematics when analyzing chemical change?

How is a balanced chemical equation used to predict amounts of species involved in a chemical process?

Big Picture

Have you ever had the opportunity to watch a launch of the Space Shuttle in person? The ability to launch rockets that leave Earth's surface and travel into space is one of humankind's greatest achievements.

The video below provides a close-up of a Space Shuttle launch from a unique perspective. This video comes from cameras located at various positions on the solid rocket booster engines, which are the long white cylinders that are located on each side of the large, orange main fuel tank. The table that appears below the video provides more information about what you are seeing in the video.

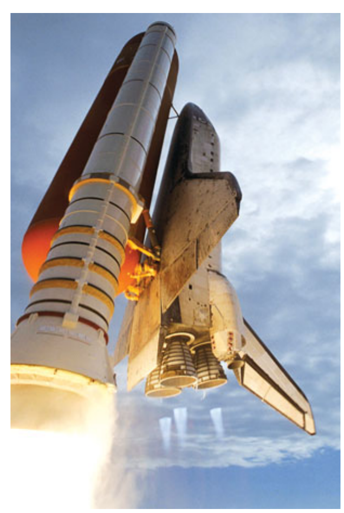

Photo courtesy NASA/JSC

Time on Video (minutes)	Event	Camera Position on Solid Rocket Booster Engine
0:00 to 1:00	Separation of solid rocket boosters and re-entry	Bottom of left-side solid rocket booster looking upward
1:01 to 2:48	Launch and separation	Top of left-side solid rocket booster looking downward
2:49 to 3:40	Separation of solid rocket boosters and re-entry	Bottom of right-side solid rocket booster looking upward
3:41 to 4:50	Launch and separation	Top of right-side solid rocket booster looking downward
4:50 to 6:00	Separation of solid rocket engines	Midway along solid rocket booster engines

Extremely vigorous chemical reactions take place in the rocket engines of the Space Shuttle. The solid rocket booster engines use one type of fuel, whereas the reaction that occurs in the main engines uses hydrogen and oxygen gases, which are contained in the large orange tank attached to the shuttle. How much hydrogen and oxygen is necessary to propel the Space Shuttle into space? Is a similar amount of each gas required? How can the balanced chemical equation (shown below) of the reaction that occurs in the main engines provide you with information about the amount of each gas required to perform this operation?

$$2\ H_2(l) + O_2(l) \rightarrow 2\ H_2O(g)$$

Examining chemical reactions using stoichiometry—an approach that considers the amounts of substances involved in a reaction—will be helpful. In order to learn about stoichiometry you will need to remind yourself of how to write, balance, and classify chemical reactions so that you can use them to interpret a chemical system. In this module you will learn to calculate the amounts of reactants and products involved in a chemical process, along with using calculations of percent yield to analyze the results of experiments you carry out.

By the end of this module you will have a better appreciation for stoichiometry. Whether you are planning to launch a rocket into space and support humans during spaceflight, or to experiment with a recipe when preparing food, a knowledge of stoichiometry is essential.

In This Module

Lesson 1—Chemical Equations

The focus of this lesson is to review the main types of chemical reactions and how to write balanced chemical equations. You will learn about what balancing means in terms of amounts of substances involved in a chemical reaction. The concepts covered in this lesson are critical to what you will learn in the remaining lessons.

- What information about a chemical system is contained within a balanced chemical equation?
- What information about a chemical system is not contained within a balanced chemical equation?
- What is conservation of mass, and how is it demonstrated in a chemical equation?

Lesson 2—Gravimetric Stoichiometry

Stoichiometry is a method to predict the amount of each substance that will be involved in a reaction. Are predictions made using this method verified by experimental results? In this lesson you will practise your skills and analyze data collected from an experiment. You will also discover how to evaluate a chemical reaction using both the theoretical and actual yield of a precipitate produced.

- How are predictions made about the masses of reactants and/or products involved in chemical reactions?
- How can you test the predictions made using the stoichiometric method?
- What is percent yield and how can it be determined?

Lesson 3—Gas Stoichiometry

In this lesson you will explore stoichiometry of chemical systems that involve the reactions of gases.

- How is the stoichiometric method applied to reactions that involve gases?

Lesson 4—Solution Stoichiometry

Do principles of stoichiometry apply to aqueous solutions? Are predictions made using this method verified by experimental results? In this lesson you will practise your skills and analyze data collected from experiments.

- How is the stoichiometric method applied to reactions that involve solutions?

Module Assessment

The assessment in this module consists of the following:

- Module 6: Lesson 1 Assignment
- Module 6: Lesson 2 Assignment
- Module 6: Lesson 3 Assignment
- Module 6: Lesson 4 Assignment

Lesson 1—Chemical Equations

Get Focused

Not all chemical reactions propel rockets into space. In most cases, nergy needed to cook food comes from a combustion reaction. A combustion reaction is just one type of chemical reaction that you may recall from your previous science studies. What other types of reactions can you recall?

The focus of this lesson is to review how to write balanced chemical equations and to discover what balancing means in terms of the amounts of substances involved in a chemical process. The concepts covered in this lesson are critical to what you will learn in the remaining lessons of this module.

Essential Questions

- What information about a chemical system is contained within a balanced chemical equation?
- What information about a chemical system is not contained within a balanced chemical equation?
- What is conservation of mass, and how is it demonstrated in a chemical equation?

Module 6: Lesson 1 Assignment

You will complete the Module 6 Assignment 1 in this lesson.

Remember that the questions that are not marked by the teacher provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.

Explore

Read

Read “2.5 Classifying Chemical Reactions” on pages 58 and 59 of your textbook, and then turn to pages 62 and 63 and read about single replacement and double replacement reactions. To summarize what you have read, copy into your notebook the information in “Table 3: Reaction Generalizations” on page 63. Place a copy of your summary in your course folder so that it is convenient for you to access it.

Self-Check

SC 1. For each of the reactants listed in question 9 on pages 66 and 67 of your textbook, identify the type of reaction that would occur. Justify your classification.

Check your work.

Self-Check Answers

SC 1.

	Reaction Type	Justification
(a)	simple decomposition	Only one reactant, a compound, is undergoing change.
(b)	formation	Both reactants are elements.
(c)	complete combustion	One of the reactants is oxygen.
(d)	double replacement	Both reactants are ionic compounds.
(e)	single replacement	One reactant is an element and the other is an ionic compound.
(f)	complete combustion	One of the reactants is oxygen.
(g)	simple decomposition	Only one reactant is shown.
(h)	formation	Both reactants are elements.
(i)	double replacement	Both reactants are ionic compounds.
(j)	double replacement	Both reactants are ionic compounds.

Read

Understanding the different types of chemical reactions is an important step in becoming able to predict the changes that may occur when reactants are combined. Review the products for each type of reaction listed in "Table 3 Reaction Generalizations" on page 63 of your textbook.

Predicting the products of a chemical reaction requires knowledge of the different types of chemical reactions as well as knowledge of how to write chemical formulas for the products.

The information below summarizes some of major aspects for writing formulas for chemical compounds. You may wish to add this information to the summary of types of chemical reactions you completed earlier in this lesson.

Some things to be aware of when writing chemical formulas are as follows:

- Some elements exist as molecular elements (e.g., H_2, O_2, Cl_2, and P_4).
- The net charge of a compound must be zero.
- Some metal ions can have more than one charge.
- The state of any ionic compound in an aqueous system is (s) or (aq).

Review your understanding of these aspects by reading the following pages in your textbook:

- "Names and Formulas of Ionic Compounds" on pages 29 to 31
- "Molecular Elements" on page 33
- how to use a solubility chart to determine the solubility of ionic compounds in aqueous systems and "Table 1: Solubility of Ionic Compounds at SATP–Generalizations," both on page 61

Self-Check

SC 2. For each of the reactants listed in question 9 on pages 66 and 67 of your textbook, predict the products of the reaction by writing their chemical formulas and respective names.

Check your work.

Self-Check Answers

SC 2.

	Chemical Formula(s) of Product(s)	Name(s) of Products
(a)	$K(s) + Cl_2(g)$	solid potassium and chlorine gas
(b)	$CuCl_2(s)$	solid copper(II) chloride (Note: The 2+ charge is the most common charge for copper.)
(c)	$CO_2(g) + H_2O(g)$	carbon dioxide gas and water vapour
(d)	$AgCl(s) + NaNO_3(aq)$	solid silver chloride and aqueous sodium nitrate
(e)	$Al(NO_3)_3(aq) + Cu(s)$	aqueous aluminium nitrate and solid copper
(f)	$CO_2(g) + H_2O(g)$	carbon dioxide gas and water vapour

(g)	$Al(s) + O_2(g)$	solid aluminium and oxygen gas
(h)	$FeBr_3(aq)$	aqueous iron(III) bromide (Note: The 3+ charge is the most common charge for iron.)
(i)	$Cu(OH)_2(s) + NaNO_3(aq)$	solid copper(II) hydroxide and aqueous sodium nitrate
(j)	$H_2O(l) + Ca_3(PO_4)_2(s)$	liquid water and solid calcium phosphate

Read

When you wrote chemical equations earlier in this course, you always ensured they were balanced. What does "balancing" an equation really mean? You may answer this question by stating that balancing an equation makes the amount of each kind of atom equal with respect to the reactants and products sides of the equation.

You may recall that balancing a chemical equation has a deeper significance—it describes how mass is conserved within a chemical system. **The law of conservation of mass states** that mass cannot be created nor destroyed. The atoms that exist today are essentially the same as those that existed millions of years ago; and over time, these atoms are recycled, breaking and forming bonds each time they react. The amount of matter at the beginning of a chemical reaction must be the same as the amount of matter at the end of the chemical reaction—all atoms and mass are accounted for.

law of conservation of mass: a law stating that in any physical or chemical system, the initial mass of the system will be identical to the final mass of the system

To refresh your memory of how to balance chemical equations, read "Balancing Chemical Equations," "Summary", and work through "Sample Problem 2.1" on pages 52 and 53. Then work through the examples below.

Example 1: Ammonia is produced by the combination of nitrogen and hydrogen gases. Balance the following chemical reaction for the formation of ammonia.

___ $N_2(g)$ + ___ $H_2(g) \rightarrow$ ___ $NH_3(g)$

Hydrogen is the atom with the largest subscript. Add a coefficient to the substances containing hydrogen on each side of the equation to balance the hydrogen atoms.

__ $N_2(g)$ + _**3**_ $H_2(g) \rightarrow$ _**2**_ $NH_3(g)$

Reactants	Products
3 × 2 = 6 H atoms	2 × 3 = 6 H atoms

Confirm that nitrogen atoms are balanced.

Reactants	Products

1 × 2 = 2 N atoms	2 × 1 = 2 N atoms

Recall that coefficients of 1 are not normally shown. Therefore, the balanced chemical equation is

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

Example 2: Ethane is one component of natural gas. Balance the complete combustion reaction for ethane.

$$\_\_\_\ C_2H_6(g) + \_\_\_\ O_2(g) \rightarrow \_\_\_\ CO_2(g) + \_\_\_\ H_2O(g)$$

Hydrogen is the atom with the largest subscript. Add a coefficient to balance the hydrogen atoms.

$$\_\_\_\ C_2H_6(g) + \_\_\_\ O_2(g) \rightarrow \_\_\_\ CO_2(g) + \_\mathbf{3}\_\ H_2O(g)$$

Add a coefficient to balance the carbon atoms.

$$\_\_\_\ C_2H_6(g) + \_\_\_\ O_2(g) \rightarrow \_\mathbf{2}\_\ CO_2(g) + \_3\_\ H_2O(g)$$

Add a coefficient to balance the oxygen atoms.

$$\_\_\_\ C_2H_6(g) + \_\mathbf{3.5}\_\ O_2(g) \rightarrow \_2\_\ CO_2(g) + \_3\_\ H_2O(g)$$

Recall that coefficients of 1 are not normally shown. Therefore, the balanced chemical equation is

$$C_2H_6(g) + 3.5\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$$

Note: If whole-number coefficients suit you better, multiply all coefficients by 2.

$$2\ C_2H_6(g) + 7\ O_2(g) \rightarrow 4\ CO_2(g) + 6\ H_2O(g)$$

Example 3: Solid silver forms as a precipitate when solid copper is placed in a solution of aqueous silver nitrate. Balance the reaction for this process.

$$\_\_\_\ Cu(s) + \_\_\_\ AgNO_3(aq) \rightarrow \_\_\_\ Ag(s) + \_\_\_\ Cu(NO_3)_2(aq)$$

Treat polyatomic ions like a single unit. The nitrate ion is the group with the largest subscript. Add a coefficient to balance the atoms in the nitrate ions.

$$\_\_\_\ Cu(s) + \_\mathbf{2}\_\ AgNO_3(aq) \rightarrow \_\_\_\ Ag(s) + \_\_\_\ Cu(NO_3)_2(aq)$$

Add a coefficient to balance the silver atoms.

$$\_\_\_\ Cu(s) + \_2\_\ AgNO_3(aq) \rightarrow \_\mathbf{2}\_\ Ag(s) + \_\_\_\ Cu(NO_3)_2(aq)$$

The copper is already balanced. Therefore, the balanced chemical equation is

$$Cu(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Cu(NO_3)_2(aq)$$

Self-Check

SC 3. Referring to your answers to question SC 2, write a balanced chemical equation for each reaction in question 9 on page 67 of the textbook.

Check your work.

Self-Check Answers

SC 3.

a) $2\ KCl(s) \rightarrow 2\ K(s) + Cl_2(g)$
b) $Cu(s) + Cl_2(g) \rightarrow CuCl_2(s)$
c) $C_5H_{12}(g) + 8\ O_2(g) \rightarrow 5\ CO_2(g) + 6\ H_2O(g)$
d) $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
e) $2\ Al(s) + 3\ Cu(NO_3)_2(aq) \rightarrow 2\ Al(NO_3)_3(aq) + 3\ Cu(s)$
f) $C_8H_{18}(g) + 12.5\ O_2(g) \rightarrow 8\ CO_2(g) + 9\ H_2O(g)$
g) $2\ Al_2O_3(s) \rightarrow 4\ Al(s) + 3\ O_2(g)$
h) $2\ Fe(s) + 3\ Br_2(l) \rightarrow 2\ FeBr_3(aq)$
i) $Cu(NO_3)_2(aq) + 2\ NaOH(aq) \rightarrow Cu(OH)_2(s) + 2\ NaNO_3(aq)$
j) $2\ H_3PO_4(aq) + 3\ Ca(OH)_2(aq) \rightarrow 6\ H_2O(l) + Ca_3(PO_4)_2(s)$

Read

When you first learned to balance equations, you may have used a table similar to the following to confirm your work.

Reaction	$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$	
Atom	Number of Atoms Before Reaction	Number of Atoms After Reaction
H	4 (2 molecules of H_2, 2 × 2H = 4)	4 (2 molecules of H_2O, 2 × 2H = 4)
O	2 (1 molecule of O_2, 1 × 2O = 2)	2 (2 molecules of H_2O, 2 × O = 2)

Read "2.3 Balancing Chemical Reaction Equations" on pages 51 to 53 of your textbook. What other ways can the information in a balanced chemical equation be interpreted? As you can see in the table, the coefficients used to balance a chemical equation can refer to the number (or amount) of each type of particle. Coefficients also refer to chemical amounts, the number of moles—the amount used to describe the chemical amount of a substance involved in a reaction.

One of the most important relationships conveyed in a balanced chemical equation is mole ratio. **Mole ratio** is a mathematical statement of the proportion of each substance involved in a chemical process relative to one another.

mole ratio: a mathematical statement of the proportion of each substance involved in a chemical process relative to one another

stoichiometry: a method of predicting or analyzing the amounts of the reactants and products participating in a chemical process

For example, in the reaction 2 $H_2(g)$ + $O_2(g)$ → 2 $H_2O(g)$, the mole ratio of hydrogen to oxygen is 2:1. This ratio shows that twice as many moles of hydrogen than oxygen are needed during the reaction. In terms of the amount (moles) of hydrogen and oxygen needed to launch the Space Shuttle, failing to place enough hydrogen into the storage tank could prevent the Space Shuttle from reaching orbit.

This example identifies why knowledge of the proportions of substances involved in a chemical process, or **stoichiometry**, is important. Stoichiometry is a method of predicting or analyzing the amounts of the reactants and products participating in a chemical process. On the Flash Drive Course or online course, play the following sound file to hear how this word is pronounced.

Mole ratios that exist in this reaction include the following:

2 $H_2(g)$ + $O_2(g)$ → 2 $H_2O(g)$	
Relationship	Mole Ratio
hydrogen to oxygen	2:1
hydrogen to water vapour	2:2 (or 1:1)
oxygen to water vapour	1:2

Example 4: If 100 mol of oxygen were to be used in a test rocket engine, how could you use mole ratio to predict the amounts of hydrogen and water vapour involved in the reaction?

Mole ratios from the balanced chemical equation will be used to predict the amount of hydrogen and oxygen involved when 100 mol of oxygen $\left(n_{O_2} = 100\text{ mol}\right)$ is reacted.

Mole ratio of

- $\text{hydrogen to oxygen} = 2:1 = \frac{2}{1}$
- $\text{water vapour to oxygen} = 2:1 = \frac{2}{1}$

Predict the amounts of hydrogen gas and water vapour.

$$n_{H_2} = n_{O_2} \times \frac{2}{1} = 100\ \text{mol} \times \frac{2}{1} = 200\ \text{mol}$$

$$n_{H_2O} = n_{O_2} \times \frac{2}{1} = 100\ \text{mol} \times \frac{2}{1} = 200\ \text{mol}$$

The predicted amounts of hydrogen and water vapour are each 200 mol.

Self-Check

SC 4. The dehydration of sugar can be represented by the following balanced chemical equation.

$$C_{12}H_{22}O_{11}(s) \rightarrow 12\ C(s) + 11\ H_2O(g)$$

Predict the amount of carbon and water vapour produced by the dehydration of 15.0 mol of sucrose, $C_{12}H_{22}O_{11}(s)$.

SC 5. Predict the amount of carbon dioxide and water vapour produced by the combustion of 398 mol of octane, $C_8H_{18}(l)$. (This is approximately the amount of gasoline in an automobile fuel tank.)

$$C_8H_{18}(g) + 12.5\ O_2(g) \rightarrow 8\ CO_2(g) + 9\ H_2O(g)$$

SC 6. Estimate the amount of ammonium sulfide required to produce a solution containing 0.253 mol of dissociated sulfide ions.

$$(NH_4)_2S(aq) \rightarrow 2\ NH_4^+(aq) + S^{2-}(aq)$$

Check your work

Self-Check Answers

SC 4. $C_{12}H_{22}O_{11}(s) \rightarrow 12\ C(s) + 11\ H_2O(g)$

List the known amounts and the mole ratios required.

$$n_{sucrose} = 15.0\ \text{mol}$$
$$n_C : n_{sucrose} = 12 : 1 = \frac{12}{1}$$
$$n_{H_2O} : n_{sucrose} = 11 : 1 = \frac{11}{1}$$

Predict the amounts of solid carbon and water vapour.

$$\begin{aligned} n_{H_2O} &= n_{sucrose} \times \frac{11}{1} \\ &= 15.0\ \text{mol} \times \frac{11}{1} \\ &= 165\ \text{mol} \end{aligned} \qquad \begin{aligned} n_{C} &= n_{sucrose} \times \frac{12}{1} \\ &= 15.0\ \text{mol} \times \frac{12}{1} \\ &= 180\ \text{mol} \end{aligned}$$

The predicted amount of solid carbon and water vapour are 180 mol and 165 mol, respectively.

SC 5. $C_8H_{18}(g) + 12.5\ O_2(g) \rightarrow 8\ CO_2(g) + 9\ H_2O(g)$

$$n_{H_2O} : n_{octane} = 9:1 = \frac{9}{1} \qquad n_{octane} = 398\ \text{mol} \qquad n_{CO_2} : n_{octane} = 8:1 = \frac{8}{1}$$

Use the mole ratios to predict the desired amounts.

$$\begin{aligned} n_{H_2O} &= n_{octane} \times \frac{9}{1} \\ &= 398\ \text{mol} \times \frac{9}{1} \\ &= 3.58 \times 10^3\ \text{mol} \end{aligned} \qquad \begin{aligned} n_{CO_2} &= n_{octane} \times \frac{8}{1} \\ &= 398\ \text{mol} \times \frac{8}{1} \\ &= 3.18 \times 10^3\ \text{mol} \end{aligned}$$

The predicted amount of carbon dioxide and water vapour are 3.18×10^3 mol and 3.58×10^3 mol, respectively

SC 6. $(NH_4)_2S(aq) \rightarrow 2\ NH_4^+(aq) + S^{2-}(aq)$

$$n_{(NH_4)_2S} : n_{S^{2-}} = 1:1 \quad n_{S^{2-}} = 0.253\ \text{mol}$$

Since the mole ratio is 1:1, the predicted amount of ammonium sulfide is 0.253 mol.

Read

Examining the quantitative aspects of a chemical reaction can provide an interesting insight. For example, can you identify the significance of the large amounts of carbon dioxide produced by the combustion of octane in question SC 5? Why does there appear to be an increase to the number of moles of the carbon compounds in this reaction? As you will see in the remaining lessons of this module, quantitative relationships are very significant when they relate to technologies that rely on chemical reactions.

Thus far, you have discovered that a great deal of information is represented within a balanced chemical equation. What kind of information does a balanced chemical equation not provide? What assumptions are made when writing or discussing a process represented by a balanced chemical equation?

Read "Chemical Reaction Equations" on pages 278 to 281 of your textbook.

Read

In Module 4, Lesson 1, you learned the effect water can have on dissolving ionic compounds. The process of dissociation creates separated ions, each with the possibility to collide with other ions and participate in a chemical reaction. To communicate which ions are participating in a chemical reaction, a unique type of balanced chemical equation is used—the net ionic equation. A net ionic equation is a type of balanced chemical equation that lists only the reacting particles.

Read "Net Ionic Equations" on pages 281 to 283 of your textbook. Work through the sample problem and examples.

Self-Check

SC 7. Complete "Practice" questions 10, 11, and 13 on page 284 of your textbook.

Check your work

Self-Check Answers

SC 7. "Practice" questions 10, 11, and 13, page 284

10. Write the balanced chemical equation.

    $Pb(s) + 2\ AgNO_3^{-}(aq) \rightarrow Pb(NO_3)_2(aq) + 2\ Ag(s)$

    Now, write the total ionic equation.

    $Pb(s) + 2\ Ag^{+}(aq) + 2\ NO_3^{-}(aq) \rightarrow Pb^{2+}(aq) + 2\ NO_3^{-}(aq) + 2\ Ag(s)$

    Cancel the ions that appear on both sides of the equation.

    $$Pb(s) + 2\ Ag^{+}(aq) + \cancel{2\ NO^{3-}(aq)} \rightarrow Pb^{2+}(aq) + \cancel{2\ NO^{3-}(aq)} + 2\ Ag(s)$$

    Therefore, the net ionic equation is

    $Pb(s) + 2\ Ag^{+}(aq) \rightarrow Pb^{2+}(aq) + 2\ Ag(s)$

11. (a)
    Balanced chemical equation: $3\ CaCl_2(aq) + 2\ Na_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + 6\ NaCl(aq)$
    Total ionic equation: $$3\ Ca^{2+}(aq) + \cancel{2\ Cl^{-}(aq)} + \cancel{6\ Na^{+}(aq)} + 2\ PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s) + \cancel{6\ Na^{+}(aq)} + \cancel{2\ Cl^{-}(aq)}$$

    Net ionic equation: $$3\ Ca^{2+}(aq) + 2\ PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$$

b) The spectator ions are chloride and sodium.

13. (a)
Balanced chemical equation: $HNO_3(aq) + NaHCO_3(aq) \rightarrow NaNO_3(aq) + H_2CO_3(aq)$

Total ionic equation:

$$H^+(aq) + \cancel{NO_3^-(aq)} + \cancel{Na^+(aq)} + HCO_3^-(aq) \rightarrow \cancel{Na^+(aq)} + \cancel{NO_3^-(aq)} + H_2CO_3(aq)$$

Net ionic equation: $H^+(aq) + HCO_3^-(aq) \rightarrow H_2CO_3(aq)$
(b) The spectator ions are nitrate and sodium.

Module 6: Lesson 1 Assignment

To complete your assignment as an online quiz, click on Module 6 Lesson 1 Assignment in the "Quizzes".

To complete your assignment as an MSWord document, select Print Course on the Flash Drive, and then **Module 6 Assignment 1**.

Reflect on the Big Picture

How does your understanding of stoichiometry help you understand a chemical system with an intended purpose? Earlier in this lesson you learned that the amounts of reactants involved in a chemical reaction can influence the amounts of products. You also learned that knowledge of the ratio by which substances react is important in the design of a reaction.

How is an understanding of stoichiometry used in cooking? In Module 5, Lesson 2, you experimented with baking soda, and you found that it would produce carbon dioxide gas with the addition of water.

Another commonly used baking product is baking powder, which contains sodium hydrogen carbonate and tartaric acid mixed together as solids. In what proportions must the components be present in order for baking powder to achieve its desired purpose? How does the design of baking powder represent an understanding of stoichiometry and provide convenience for people using the product?

Save your answers in your course folder.

You may wish to share your answers with your classmates or some other people and make revisions based on their comments.

Assessment

- Unit D Module 6 Assignment 1

Lesson Summary

In this lesson you explored the following questions:

- What information about a chemical system is contained within a balanced chemical equation?
- What information about a chemical system is not contained within a balanced chemical equation?
- What is conservation of mass, and how is it demonstrated in a chemical equation?

You were introduced to stoichiometry, a mathematical process to determine amounts of substances involved in a chemical process. You learned that the information contained within balanced chemical equations provides information about chemical amounts, such as the number of moles of substances involved.

You discovered that the coefficients of a balanced chemical equation describe the proportions of each substance involved in the reaction, and how these proportions are an essential part of the predictions made using stoichiometry.

You learned that analyzing a system in terms of number of moles of substances does not allow you to comment on measurable amounts like mass or volume. In the lessons to come, you will extend your understanding of stoichiometry to use measurements of matter to predict the outcomes of a variety of chemical reactions.

Lesson Glossary

law of conservation of mass: a law stating that in any physical or chemical system, the initial mass of the system will be identical to the final mass of the system

mole ratio: a mathematical statement of the proportion of each substance involved in a chemical process relative to one another

net ionic equation: a type of balanced chemical equation that lists only the reacting particles

stoichiometry: a method of predicting or analyzing the amounts of the reactants and products participating in a chemical process

Lesson 2—Gravimetric Stoichiometry

Get Focused

In Lesson 1 you learned that stoichiometry is a method to predict the amounts of reactants or products involved in a chemical reaction. Whether you plan to launch a rocket, and need to calculate oxygen to fuel ratios, or need to calculate the amount of product that will result from a chemical reaction, a knowledge of stoichiometry is essential. In Lesson 1 you concentrated on the significance of ratios among the coefficients of different substances in the reaction to this method of calculation.

There are two major drawbacks to the predictions you have made thus far. First, you have only been able to consider chemical amounts of each substance (number of moles) involved. You already know that the number of moles is not a directly measurable amount. Second, as is the case with all predictions, they need to be confirmed.s

In this lesson you will continue your study of stoichiometry by learning how to use the masses of the substances involved in your predictions. You will also perform an experiment to test predictions made using stoichiometry. By completing these activities, you will perform a gravimetric analysis. **Gravimetric analyses** involve the determination of masses of substances. Open the following sound (either on the Flash Drive course or access it online) file to hear how this word is pronounced.

gravimetric analysis: the determination of masses of substances

Essential Questions

- How are predictions made about the masses of reactants and/or products involved in chemical reactions?
- How can you test the predictions made using the stoichiometric method?
- What is percent yield and how can it be determined?

Module 6: Lesson 2 Assignment

You will complete the Module 6 Assignment 2 in this lesson.

Remember that the questions that are not marked by the teacher provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.

Explore

Read

In Lesson 1 of this module you examined the chemical reaction that occurs in the main engines of the Space Shuttle.

$$2\ H_2(l) + O_2(l) \rightarrow 2\ H_2O(g)$$

From your work in Lesson 1, you identified important aspects of this balanced chemical equation, such as the mole ratio for hydrogen and oxygen. You would correctly interpret this ratio as meaning that twice as many moles of hydrogen than oxygen are needed for this reaction.

What would happen if a model for this reaction were to be planned using 2.0 g of hydrogen and 1.0 g of oxygen? Would the proper stoichiometric ratio for the reaction be met? Read and follow the calculations given in the table to answer this question.

Reactant	$H_2(l)$	$O_2(l)$
Mass	2.0 g	1.0 g
Molar Mass	2.02 g/mol	32.00 g/mol
Number of Moles $\left(n = \frac{m}{M}\right)$	0.99 mol	0.031 mol
Mole Ratio of (H_2:O_2)	32:1	
Desired Mole Ratio of (H_2:O_2)*	2:1	
* The desired mole ratio is derived from the balanced chemical equation.		

Compare the mole ratio calculated from the masses of each reactant with the desired mole ratio from the balanced chemical equation. Would the correct proportion of hydrogen be used in this model reaction? Can you provide an explanation as to why the correct proportion of these two reactants could not be met?

As you may have identified in the previous example, hydrogen and oxygen have significantly different molar masses. This was not accounted for when suggesting masses of each gas to use in the model reaction. As a result, there are very different numbers of moles for the two reactants. What effect would completing this reaction with these masses have?

If the reaction were allowed to occur with the masses suggested, a large amount of hydrogen would not have reacted, simply because not enough oxygen was present. In Module 7 you will learn more about how using amounts of reactants that are not stoichiometric can affect the yield of a reaction. The yield of a reaction is a measured amount of product obtained by a chemical reaction. It is often expressed as a percentage of maximum yield. You may have also thought that since all the hydrogen was not able to react, a much reduced amount of energy would have been produced by the reaction as well.

yield of a reaction: a measured amount of product obtained by a chemical reaction often expressed as a percentage of maximum yield.

What mass of hydrogen would be required to react with all the oxygen in the model reaction? How could you use your knowledge of stoichiometry to calculate this?

By examining this example, you have seen that making predictions using stoichiometry must always focus on the number of moles of each substance involved. Unfortunately, there is no device to measure the number of moles of each substance. Therefore, to complete predictions using stoichiometry when provided the mass of a substance involved, convert the mass into number of moles using this relationship:

$$n = \frac{m}{M}$$

Work through the following example to see how stoichiometric predictions can be made when you are provided with the mass of one reactant in a chemical reaction.

Example 1: For testing a rocket engine, what mass of hydrogen gas is required to completely react with 1.0 g of oxygen gas?

Step 1: Write the balanced chemical equation for the reaction.

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$$

Step 2: Organize the information provided.

2 H_2(g)	**+**	**O_2(g)**	**→**	**2 H_2O(g)**
$m = ?$ $M_{H_2} = 2(M \text{ of H})$ $= 2(1.01 \text{ g/mol})$ $= 2.02 \text{ g/mol}$		$m = 1.0 \text{ g}$ $M_{O_2} = 2(M \text{ of O})$ $= 2(16.00 \text{ g/mol})$ $= 32.00 \text{ g/mol}$		

Step 3: Calculate the number of moles of oxygen.

$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{1.0 \text{ g}}{32.00 \text{ g/mol}} \\ &= 0.031\,25 \text{ mol} \end{aligned}$$

Step 4: Use the mole ratio between hydrogen and oxygen to determine the number of moles of hydrogen.

$$\begin{aligned} n_{H_2} &= n_{O_2} \times \frac{2_{H_2}}{1_{O_2}} \\ &= 0.031\,25 \text{ mol} \times \frac{2_{H_2}}{1_{O_2}} \\ &= 0.0625 \text{ mol} \end{aligned}$$

Step 5: Calculate the mass of hydrogen required.

$$
\begin{aligned}
n &= \frac{m}{M} \\
m &= nM \\
&= (0.0625\ \text{mol})(2.02\ \text{g/mol}) \\
&= 0.13\ \text{g}
\end{aligned}
$$

To completely react 1.0 g of oxygen gas, 0.13 g of hydrogen gas are required.

Read "Calculating Masses Involved in Chemical Reactions" on page 288 of your textbook. Then work through the "SAMPLE problem 7.2" and "COMMUNICATION example" on pages 288 and 289.

Self-Check

SC 1. Complete "Practice" questions 9 to 14 on page 290 of your textbook.

Check your work.

Self-Check Answers

SC 1. Practice questions 9 to 14, page 290

9.

8 Zn(s)	**+**	**S_8(s)**	**→**	**8 ZnS(s)**
$m = 25$ g $M_{Zn} = 65.41$ g/mol		$m = ?$ $M_{S_8} = 8(M \text{ of S})$ $= 8(32.07\ \text{g/mol})$ $= 256.56\ \text{g/mol}$		

Determine the number of moles of zinc.

$$
\begin{aligned}
n &= \frac{m}{M} \\
&= \frac{25\ \text{g}}{65.41\ \text{g/mol}} \\
&= 0.382\ 204\ 555\ 9\ \text{mol}
\end{aligned}
$$

Use the mole ratio to determine the number of moles of sulfur.

$$
\begin{aligned}
n_{S_8} &= n_{Zn} \times \frac{1}{8} \\
&= 0.382\ 204\ 555\ 9\ \text{mol} \times \frac{1}{8} \\
&= 0.047\ 775\ 569\ 5\ \text{mol}
\end{aligned}
$$

Calculate the mass of sulfur required.

$$n = \frac{m}{M}$$

$$\begin{aligned} m &= nM \\ &= (0.047\ 775\ 569\ 5\ \text{mol})(256.56\ \text{g/mol}) \\ &= 12\ \text{g} \end{aligned}$$

12 g of sulfur is required to react with 25 g of zinc.

10.

2 $Al_2O_3(s)$	→	**2 $Al(s)$**	+	**3 $O_2(g)$**
$m = 125\ \text{g}$ $M_{Al_2O_3} = 2(M \text{ of Al}) + 3(M \text{ of O})$ $= 2(26.98\ \text{g/mol}) + 3(16.00\ \text{g/mol})$ $= 101.96\ \text{g/mol}$		$m = ?$ $M_{Al} = 26.98\ \text{g/mol}$		

$$n_{Al} = \frac{m}{M}$$

$$\begin{aligned} m &= n_{Al}M \\ &= (1.225\ 970\ 969\ \text{mol})(26.98\ \text{g/mol}) \\ &= 33.1\ \text{g} \end{aligned}$$

$$\begin{aligned} n_{Al_2O_3} &= \frac{m}{M} \\ &= \frac{125\ \text{g}}{101.96\ \text{g/mol}} \\ &= 1.225\ 970\ 969\ \text{mol} \end{aligned}$$

$$\begin{aligned} n_{Al} &= n_{Al_2O_3} \times \frac{2}{2} \\ &= 1.225\ 970\ 969\ \text{mol} \times \frac{2}{2} \\ &= 1.225\ 970\ 969\ \text{mol} \end{aligned}$$

33.1 g of aluminium is produced when 125 g of bauxite is decomposed.

$C_3H_8(g)$	+	$5\ O_2(g)$	→	$3\ CO_2(g)$	+	$4\ H_2O(g)$
$m = ?$ $M_{O_2} = 2(M \text{ of O}) = 2(16.00 \text{ g/mol}) = 32.00 \text{ g/mol}$		$m = 10.0 \text{ g}$ $M_{C_3H_8} = 3(M \text{ of C}) + 8(M \text{ of H}) = 3(12.01 \text{ g/mol}) + 8(1.01 \text{ g/mol}) = 44.11 \text{ g/mol}$				

$$
\begin{aligned}
n_{O_2} &= \frac{m}{M} \\
m_{O_2} &= nM \\
&= (1.133\ 529\ 812 \text{ mol})(32.00 \text{ g/mol}) \\
&= 36.3 \text{ g}
\end{aligned}
$$

$$
\begin{aligned}
n_{C_3H_8} &= \frac{m}{M} \\
&= \frac{10.0 \text{ g}}{44.11 \text{ g/mol}} \\
&= 0.226\ 705\ 962\ 4 \text{ mol}
\end{aligned}
$$

$$
\begin{aligned}
n_{O_2} &= n_{C_3H_8} \times \frac{5}{1} \\
&= 0.226\ 705\ 962\ 4 \text{ mol} \times \frac{5}{1} \\
&= 1.133\ 529\ 812 \text{ mol}
\end{aligned}
$$

36.3 g of oxygen is required to combust 10.0 g of propane.

12.

$2\ NaCl(aq)$	+	$Pb(NO_3)_2(aq)$	→	$PbCl_2(s)$	+	$2\ NaNO_3(aq)$
$m = 2.57 \text{ g}$ $M_{NaCl} = (M \text{ of Na}) + (M \text{ of Cl}) = (22.99 \text{ g/mol}) + (35.45 \text{ g/mol}) = 58.44 \text{ g/mol}$				$m = ?$ $M_{PbCl_2} = (M \text{ of Pb}) + 2(M \text{ of Cl}) = (207.2 \text{ g/mol}) + 2(35.45 \text{ g/mol}) = 278.10 \text{ g/mol}$		

9.

2 NaCl(aq)	+	$Pb(NO_3)_2$(aq)	→	$PbCl_2$(s)	+	2 $NaNO_3$ (aq)
$m = 2.57\ \text{g}$ $M_{NaCl} = (M \text{ of Na}) + (M \text{ of Cl})$ $= (22.99\ \text{g/mol}) + (35.45\ \text{g/mol})$ $= 58.44\ \text{g/mol}$				$m = ?$ $M_{PbCl_2} = (M \text{ of Pb}) + 2(M \text{ of Cl})$ $= (207.2\ \text{g/mol}) + 2(35.45\ \text{g/mol})$ $= 278.10\ \text{g/mol}$		

$$\begin{aligned} n_{NaCl} &= \frac{m}{M} \\ &= \frac{2.57\ \text{g}}{58.44\ \text{g/mol}} \\ &= 0.043\ 976\ 728\ 3\ \text{mol} \end{aligned}$$

$$\begin{aligned} n_{PbCl_2} &= n_{NaCl} \times \frac{1}{2} \\ &= 0.043\ 976\ 728\ 3\ \text{mol} \times \frac{1}{2} \\ &= 0.021\ 988\ 364\ 1\ \text{mol} \end{aligned}$$

$$\begin{aligned} n_{PbCl_2} &= \frac{m}{M} \\ m_{PbCl_2} &= nM \\ &= (0.021\ 988\ 364\ 1\ \text{mol})(278.10\ \text{g/mol}) \\ &= 6.11\ \text{g} \end{aligned}$$

6.11 g of lead(II) chloride precipitate is produced by the reaction.

10.

2 Al(s)	+	3 H_2SO_4(aq)	→	3 H_2(g)	+	$Al_2(SO_4)_3$(aq)
$m = 2.73\ \text{g}$ $M_{Al} = 26.98\ \text{g/mol}$				$m = ?$ $M_{H_2} = 2(M \text{ of H})$ $= 2(1.01\ \text{g/mol})$ $= 2.01\ \text{g/mol}$		

11.

$$\begin{aligned} n_{Al} &= \frac{m}{M} \\ &= \frac{2.73\ g}{26.98\ g/mol} \\ &= 0.101\ 186\ 063\ 8\ mol \end{aligned}$$

12.

$$\begin{aligned} n_{H_2} &= n_{Al} \times \frac{3}{2} \\ &= 0.101\ 186\ 063\ 8\ mol \times \frac{3}{2} \\ &= 0.151\ 779\ 095\ 6\ mol \end{aligned}$$

$$\begin{aligned} n_{H_2} &= \frac{m}{M} \\ m_{H_2} &= nM \\ &= (0.151\ 779\ 095\ 6\ mol)(2.02\ g/mol) \\ &= 0.307\ g \end{aligned}$$

0.307 g of hydrogen is produced when 2.73 g of aluminium react.

13.

2 KOH(aq)	**+**	**$Cu(NO_3)_2$(aq)**
$m = 2.67\ g$ $M_{KOH} = (M\ of\ K) + (M\ of\ O) + (M\ of\ H)$ $= (39.10\ g/mol) + (16.00\ g/mol) + (1.01\ g/mol)$ $= 56.11\ g/mol$		

→	**$Cu(OH)_2$(s)**	**+**	**2 KNO_3(aq)**
	$m = ?$ $M_{Cu(OH)_2} = (M\ of\ Cu) + 2(M\ of\ O) + 2(M\ of\ O)$ $= (63.55\ g/mol) + 2(16.00\ g/mol) + 2(1.01\ g/mol)$ $= 97.57\ g/mol$		

$$
\begin{aligned}
n_{Cu(OH)_2} &= n_{KOH} \times \frac{1}{2} \\
&= 0.047\ 585\ 100\ 7\ \text{mol} \times \frac{1}{2} \\
&= 0.023\ 792\ 550\ 3\ \text{mol}
\end{aligned}
$$

$$
\begin{aligned}
n_{KOH} &= \frac{m}{M} \\
&= \frac{2.67\ \text{g}}{56.11\ \text{g/mol}} \\
&= 0.047\ 585\ 100\ 7\ \text{mol}
\end{aligned}
$$

$$
\begin{aligned}
n_{Cu(OH)_2} &= \frac{m}{M} \\
m_{Cu(OH)_2} &= nM \\
&= (0.023\ 792\ 550\ 3\ \text{mol})(97.57\ \text{g/mol}) \\
&= 2.32\ \text{g}
\end{aligned}
$$

2.32 g of copper(II) hydroxide precipitate is produced by the reaction.

Read

Read "Testing the Stoichiometric Method" on pages 290 and 291. Then read "Evaluation" on pages 791 and 792.

Try This

Read "LAB EXERCISE 7.A: Testing the Stoichiometric Method" on page 291 of your textbook.

SC 2. Predict the mass of lead produced using the stoichiometric method.

SC 3. According to the evidence gathered, determine the actual mass of lead produced in this experiment.

SC 4. What is the percent difference between the experimental and predicted results?

Check your work.

Self-Check Answers

SC 2.

Zn(s)	+	$Pb(NO_3)_2(aq)$	→	Pb(s)	+	$Zn(NO_3)_2(aq)$
$m = 2.13$ g $M_{Zn} = 65.41$ g/mol				$m = ?$ $M_{Pb} = 207.2$ g/mol		

$$\begin{aligned} n_{Zn} &= \frac{m}{M} \\ &= \frac{2.13\text{ g}}{65.41\text{ g/mol}} \\ &= 0.032\,563\,828\,2\text{ mol} \end{aligned}$$

$$\begin{aligned} n_{Pb} &= n_{Zn} \times \frac{1}{1} \\ &= 0.032\,563\,828\,2\text{ mol} \times \frac{1}{1} \\ &= 0.032\,563\,828\,2\text{ mol} \end{aligned}$$

$$\begin{aligned} n_{Pb} &= \frac{m}{M} \\ m_{Pb} &= nM \\ &= (0.032\,563\,828\,2\text{ mol})(207.2\text{ g/mol}) \\ &= 6.75\text{ g} \end{aligned}$$

According to the stoichiometric method, 6.75 g of lead is produced by the reaction of 2.13 g of zinc.

SC 3. m = 7.60 g − 0.92 g = 6.68 g

According to the evidence gathered, 6.68 g of lead was produced.

SC 4.

$$\begin{aligned} \%\text{ difference} &= \frac{|\text{experimental value} - \text{predicted value}|}{\text{predicted value}} \times 100\% \\ &= \frac{|6.68\text{ g} - 6.75\text{ g}|}{6.75\text{ g}} \times 100\% \\ &= 1.04\% \end{aligned}$$

The experiment is judged to be effective since the percentage difference is less than 5%.

Try This

Testing any predictions made using the stoichiometric method involves performing an experiment and comparing the prediction with the measurable amount (result) obtained. In the following activities you will test the stoichiometric method using gravimetric analysis.

In the pre-lab component you will calculate the expected result for the experiment. Following the pre-lab, you will use a virtual lab investigation to perform an experiment. Part of your analysis of the investigation will require you to determine a percentage difference between the predicted and observed mass of product. For the purpose of this investigation and the equipment used, a percentage difference of less than 5% would support the hypothesis that predictions made using the stoichiometric method are supported by experimental evidence.

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Pre-lab: Decomposition of Malachite

Malachite is commonly used in jewellery. Malachite can react when heated. In this investigation you will heat malachite to test whether the predictions made using the stoichiometric method are confirmed by experimental data.

Note: The correct chemical formula for malachite is $Cu(OH)_2CuCO_3(s)$, not what is stated in your textbook on page 288. The reaction of malachite when exposed to heat is as follows:

$$Cu(OH)_2CuCO_3(s) \rightarrow 2CuO(s) + CO_2(g) + H_2O(g)$$

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Predict the mass of copper(II) oxide that will remain if 2.00 g of malachite undergoes decomposition when heated.
Record and save your prediction. It will be required in your assignment.

Read

Read "Applications of Stoichiometry" on page 292 of your textbook.

Self-Check

SC 5. Complete questions 8 and 10 of "Section 7.2 Questions" on page 293 of your textbook.

Check your work.

Self-Check Answers

SC 2. Section 7.2 Questions 8 and 10, page 293

8. a.

2 $AgNO_3$(aq)	**+**	**Na_2CrO_4(aq)**	**→**
$m = 3.00\ \text{g}$ $M_{AgNO_3} = (M \text{ of Ag}) + (M \text{ of N}) + 3(M \text{ of O})$ $= (107.87\ \text{g/mol}) + (14.01\ \text{g/mol}) + 3(16.00\ \text{g/mol})$ $= 169.88\ \text{g/mol}$			

Ag_2CrO_4(s)	**+**	**2 $NaNO_3$(aq)**	**→**
$m = ?$ $M_{Ag_2CrO_4} = 2(M \text{ of Ag}) + (M \text{ of Cr}) + 4(M \text{ of O})$ $= 2(107.87\ \text{g/mol}) + (52.00\ \text{g/mol}) + 4(16.00\ \text{g/mol})$ $= 331.74\ \text{g/mol}$			

$$\begin{aligned} n_{AgNO_3} &= \frac{m}{M} \\ &= \frac{3.00\ \text{g}}{169.88\ \text{g/mol}} \\ &= 0.017\,659\,524\,4\ \text{mol} \end{aligned}$$

$$\begin{aligned} n_{Ag_2CrO_4} &= n_{AgNO_3} \times \frac{1}{2} \\ &= 0.017\,659\,524\,4\ \text{mol} \times \frac{1}{2} \\ &= 0.008\,829\,762\,2\ \text{mol} \end{aligned}$$

$$n_{Ag_2CrO_4} = \frac{m}{M}$$
$$\begin{aligned} m_{Ag_2CrO_4} &= nM \\ &= (0.008\,829\,762\,2 \text{ mol})(331.74 \text{ g/mol}) \\ &= 2.93 \text{ g} \end{aligned}$$

The mass of the precipitate is predicted to be 2.93 g.

b.

$$\begin{aligned} \text{percent yield} &= \frac{\text{actual yield}}{\text{predicted yield}} \times 100\% \\ &= \frac{2.81 \text{ g}}{2.93 \text{ g}} \times 100\% \\ &= 95.9\% \end{aligned}$$

10.a.

$BaCl_2(aq)$	+	$Na_2SO_4(aq)$	→
$m = 9.8$ g $M_{BaCl_2} = (M \text{ of Ba}) + 2(M \text{ of Cl})$ $= (137.33 \text{ g/mol}) + 2(35.45 \text{ g/mol})$ $= 208.23$ g/mol			

$BaSO_4(s)$	+	$2\ NaCl(aq)$
$m = ?$ $M_{BaSO_4} = (M \text{ of Ba}) + (M \text{ of S}) + 4(M \text{ of O})$ $= (137.33 \text{ g/mol}) + (32.07 \text{ g/mol}) + 4(16.00 \text{ g/mol})$ $= 233.40$ g/mol		

$$\begin{aligned} n_{BaSO_4} &= n_{BaCl_2} \times \frac{1}{1} \\ &= 0.047\,063\,343\,4 \text{ mol} \times \frac{1}{1} \\ &= 0.047\,063\,343\,4 \text{ mol} \end{aligned}$$

$$n_{Ag_2CrO_4} = \frac{m}{M}$$
$$\begin{aligned} m_{BaSO_4} &= nM \\ &= (0.047\,063\,343\,4 \text{ mol})(233.40 \text{ g/mol}) \\ &= 11 \text{ g} \end{aligned}$$

$$
\begin{aligned}
n_{BaCl_2} &= \frac{m}{M} \\
&= \frac{9.8\ g}{208.23\ g/mol} \\
&= 0.047\ 063\ 343\ 4\ mol
\end{aligned}
$$

The mass of the precipitate is predicted to be 11 g.

b.

$$
\begin{aligned}
\text{percent yield} &= \frac{\text{actual yield}}{\text{predicted yield}} \times 100\% \\
&= \frac{10.0\ g}{11\ g} \times 100\% \\
&= 91\%
\end{aligned}
$$

c. The percent yield suggests that minor mistakes were made during the experiment. These mistakes may have involved a loss of product while handling the precipitate or a loss of reactant during the preparation. The result is within the 5%–10% margin normally expected for school laboratory work.

Module 6: Lesson 2 Assignment

You will complete the questions for the following lab as your assignment for this lesson.

Lab: Decomposition of Malachite

Problem

Are predictions made using the stoichiometric method supported by experimental evidence?

Purpose

The following virtual lab will test the prediction you made in the pre-lab exercise. Before beginning the lab, ensure you read through all parts of the laboratory, including the background, procedure, and assignment. As you read through the procedure, you may wish to make a data table to record your observations.

Decomposition of Malachite Virtual Lab

Module 6: Lesson 2 Assignment

To complete your assignment as an online quiz, click on Module 6 Lesson 2 Assignment in the "Quizzes".

To complete your assignment as an MSWord document, click Print Course, on the Flash Drive, and select **Module 6 Assignment 2**.

Reflect on the Big Picture

At the beginning of this lesson you examined how using appropriate amounts of hydrogen and oxygen can affect the products and usefulness of a chemical reaction. Knowledge of stoichiometry allows you to use appropriate amounts of reactants in order to maintain safety in many situations.

Assessment

- Module 6: Assignment 2

Lesson Summary

In this lesson you investigated the following questions:

- How are predictions made about the masses of reactants and/or products involved in chemical reactions?
- How can you test the predictions made using the stoichiometric method?
- What is percent yield and how can it be determined?

The procedure for calculating the masses of reactants or products in a chemical reaction is called gravimetric stoichiometry. There are four steps in solving any gravimetric stoichiometry problem:

Step 1: Write a balanced chemical equation, and list the given information—the masses and molar masses of the known and desired substances.

Step 2: Convert the mass of the known substance into number of moles.

Step 3: Use the mole ratio to convert the number of moles calculated in Step 2 into the number of moles for the desired substance.

Step 4: Calculate the mass of the desired substance using the number of moles calculated in Step 3.

You discovered that the stoichiometric method can be tested by comparing predicted values with data collected through experimentation. You also calculated the percent difference and found that percent difference is a useful way to determine the validity of predictions made using the stoichiometric method.

You also calculated percent yield. Percent yield is the ratio of the the actual experimental amount of product obtained and the maximum possible theoretical amount of product (as determined by the stoichiometry method). In Lesson 3 you will use percent yield to assess the results of an experiment.

Lesson Glossary

gravimetric analysis: **the determination of masses of substances**

yield of a reaction: a measured amount of product obtained by a chemical reaction, often expressed as a percentage of maximum yield

Lesson 3—Gas Stoichiometry

Get Focused


Photo courtesy Kennedy Space Center/NASA

The hydrogen and oxygen that power the main engines of the Space Shuttle during launch are placed into the large orange external fuel tank mounted to the shuttle (as seen in this photo). The white ring that appears at the top of the external fuel tank is used to transfer hydrogen and oxygen (as liquids) into the tank.

In Lesson 2 you discovered that stoichiometry can be used to predict the mass of each of these substances that needs to be placed into the external fuel tank. During the combustion of hydrogen and oxygen, the temperature of the main engines of the shuttle can exceed 3300°C. At these conditions, the water produced by the reaction would exist as a gas. As you may recall from your study in Module 4, gases may fluctuate in volume in response to the temperature and pressure at which they are contained.

Predicting the amount of a gas that is involved in a chemical reaction must still involve the number of moles of substance; but, in this case, the number of moles will be measured using the pressure, volume, and temperature of the gas. Can you recall a relationship that includes these four variables? How would you use this relationship when predicting the amount of a gaseous substance involved in a chemical reaction?

In this lesson you will use stoichiometry to predict the amounts of substances in chemical reactions that involve gases.

Essential Questions

- How is the stoichiometric method applied to reactions that involve gases?

Module 6: Lesson 3 Assignment

You will complete the Module 6 Assignment 3 in this lesson.

Remember that the questions that are not marked by the teacher provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.

Explore

Read

People often don't think about it, but chemical reactions involving gases are quite common. Do you recall the ideal gas law? This law was an important relationship that equated the number of moles of a gas to the conditions at which it is contained (pressure, volume, and temperature).

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Read "7.3 Gas Stoichiometry" on pages 294 to 296 of your textbook. Work through "SAMPLE problem 7.3" and "COMMUNICATION examples" 1 and 2.

Use your notes or your textbook to write the following relationships that you will refer to many times in this lesson:

- ideal gas law
- volume of a gas at standard temperature and pressure (STP)
- volume of a gas at standard ambient temperature and pressure (SATP)

Discuss


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Are the general steps for stoichiometry calculations the same for both solids and gases? Do you agree with the statements and explanation provided in the paragraph following “COMMUNICATION example” 1 on page 295 of your textbook? Would it be possible to summarize the methods used to complete stoichiometry calculations for both solids and gases? Are the “SUMMARY” and the “Stoichiometry Calculations” in the margin on page 296 of your textbook adequate?
Save a copy of your answers to your course folder. You may wish to share your answers with your classmates or some other people.

Self-Check

Complete “Practice” questions 1 to 3 on page 296 of your textbook.

Check your work.

Self-Check Answers

Practice questions 1 to 3, page 296

1.

$2\ CH_3OH(g)$	+	$3\ O_2(g)$	→
$m = 15\ \text{g}$ $M_{CH_3OH} = (M \text{ of C}) + 4(M \text{ of H}) + (M \text{ of O})$ $= (12.01\ \text{g/mol}) + 4(1.01\ \text{g/mol}) + (16.00\ \text{g/mol})$ $= 32.05\ \text{g/mol}$		$V = ?$ $\text{STP} = 22.4\ \text{L/mol}$	

$2\ CO_2(g)$	+	$4\ H_2O(g)$

$$n_{CH_3OH} = \frac{m}{M} = \frac{15\ \text{g}}{32.05\ \text{g/mol}} = 0.468\ 018\ 720\ 7\ \text{mol}$$

$$n_{O_2} = n_{CH_3OH} \times \frac{3}{2} = 0.468\ 018\ 720\ 7\ \text{mol} \times \frac{3}{2} = 0.702\ 028\ 081\ 1\ \text{mol}$$

$$V_{O_2} = n_{CH_3OH}(22.4\ \text{L/mol}) = (0.702\ 028\ 081\ 1\ \text{mol})(22.4\ \text{L/mol}) = 16\ \text{L}$$

The combustion of 15 g of methanol requires 16 L of oxygen gas at STP.

2.

2 NaCl(l)	**+**	**2 Na(s)**	**→**	**Cl_2(g)**
		$m = 105\ \cancel{\text{kg}} \times \frac{1000\ \text{g}}{1\ \cancel{\text{kg}}} = 1.05 \times 10^5\ \text{g}$ $M_{Na} = 22.99\ \text{g/mol}$		V = ? T = 30 °C + 273 = 303 K P = 95.7 kPa

$$n_{Na} = \frac{m}{M} = \frac{1.05 \times 10^5\ \text{g}}{22.99\ \text{g/mol}} = 4567.203\ 132\ \text{mol}$$

$$n_{Cl_2} = n_{Na} \times \frac{1}{2} = 4567.203\ 132\ \text{mol} \times \frac{1}{2} = 2283.601\ 566\ \text{mol}$$

$$V_{O_2} = \frac{nRT}{P} = \frac{(2283.601\ 566\ \cancel{\text{mol}})(8.314\ \cancel{\text{kPa}}\bullet\text{L}/\cancel{\text{mol}}\bullet\cancel{\text{K}})(303\ \cancel{\text{K}})}{95.7\ \cancel{\text{kPa}}} = 6.0 \times 10^4\ \text{L}$$

Along with 105 kg of sodium metal, 6.0×10^4 L of chlorine gas is produced.

3.

2 H_2(g)	**+**	**O_2(g)**	**→**	**2 H_2O(g)**
V = 300 L *T* = 40 °C + 273 = 313 K *P* = 1.50 atm		*V* = ? *T* = 40 °C + 273 = 313 K *P* = 1.50 atm		

The law of combining volumes states that the volumes of gaseous reactants and products of a chemical reaction (when measured at the same temperature and pressure) are always in simple ratios of whole numbers. Therefore,

$$\begin{aligned} V_{O_2} &= V_{H_2} \times \frac{1}{2} \\ &= 300\ \text{L} \times \frac{1}{2} \\ &= 150\ \text{L} \end{aligned}$$

Lab: Producing Hydrogen

Purpose

In Lesson 2 you tested predictions that were made using the stoichiometric method. In this lab activity you will perform another experiment to test whether predictions made using the stoichiometric method are valid for gases.

To prepare for this investigation, read “INVESTIGATION 7.3: Producing Hydrogen” on pages 305 and 306 of your textbook. Construct a data table in the space provided in your Module 6: Lesson 3 Assignment. As you read through the procedure, record your observations in the table.

Problem

Are predictions made using the stoichiometric method supported by experimental evidence?

Procedure

View a virtual presentation of this investigation, **Producing Hydrogen**.

Module 6: Lesson 3 Assignment

To do your assignment as an online quiz, click on Module 6 Lesson 2 Assignment in the "Quizzes" and answer questions 1 and 2.

To do your assignment as an MSWord document, click Print course, on the Flash Drive, then select **Module 6 assignment 3** and answer questions 1 and 2.

Case Study: Producing Hydrogen for Fuel Cells

Throughout this module you have considered the use of hydrogen and oxygen as reactants in the main engines of the Space Shuttle. Hydrogen and oxygen gases can also be used in fuel cells—devices that convert chemical potential energy into electrical energy. Hydrogen-oxygen fuel cells have been an essential component of the space program. The fuel cells have produced electricity to run the electronic components within spacecrafts since the

1960s. Many questions exist concerning possible sources of hydrogen for fuel cells. The lack of clear answers to these questions has limited the large scale use of these devices.

Read "Case Study: Producing Hydrogen for Fuel Cells" on pages 297 and 298 of your textbook.

Module 6: Lesson 3 Assignment

Complete the rest of the questions in your Module 6 Lesson 3 Assignment.

Reflect on the Big Picture


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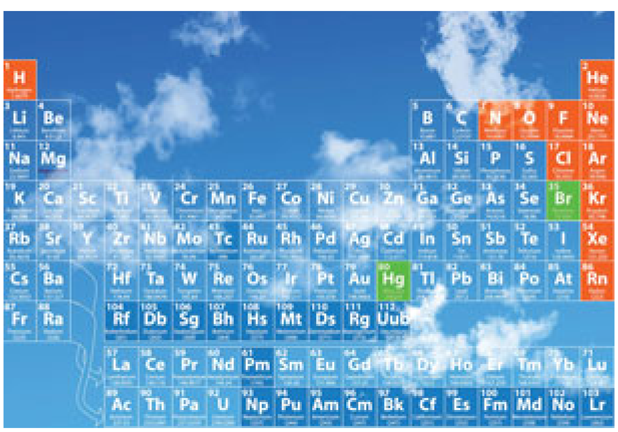
Further development of fuel-cell technology and processes to produce hydrogen may mean that cars that use fuel cells may become more common in the future. In 2007, hydrogen-powered vehicles—due in part to the stoichiometry of the energy-producing reaction—had a limited driving range. In addition, few locations existed for refilling hydrogen-powered vehicles.



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Assessment

- Module 6: Lesson 3 Assignment

Lesson Summary

In this lesson you explored the following essential question:

- How is the stoichiometric method applied to reactions that involve gases?

You learned to apply your knowledge of stoichiometry to predict the amounts of gaseous substances involved in chemical reactions. You also completed a laboratory experiment where hydrogen gas was produced in a chemical reaction. You then saw examples of how knowledge of the stoichiometry of gases is important to the operation of many technologies, including combustion processes and hydrogen-oxygen fuel cells.

In the next lesson you will look at the stoichiometry of reactions that involve solutions.

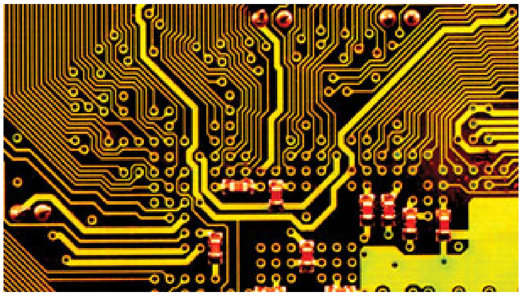
Lesson 4—Solution Stoichiometry

Get Focused

The ability to launch humans into space and then support them while they are there involves the coordination of many systems. It might surprise you that circuit boards like the one pictured below are involved in all systems of a spacecraft. Whether the system is designed to control the flow of fuel into a rocket engine or to control the composition of gases within the crew compartment of the spacecraft, electronic circuits are involved.

Image courtesy of NASA

The manufacture of a circuit board actually involves removing the copper metal from a plate. Proper treatment and handling of the industrial waste from this process attempts to reclaim the copper. The process involves knowledge of the stoichiometry of aqueous solutions.



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In the previous lessons you learned to use stoichiometry to predict quantitative changes of solid and gaseous substances. As you were shown in Modules 4 and 5, many chemical processes occur in aqueous solutions. Knowledge of stoichiometry involving solutions is essential in understanding many industrial and environmental issues.

In this lesson you will use stoichiometry to predict the amounts of substances in chemical reactions that involve solutions.

Essential Questions

- How is the stoichiometric method applied to reactions that involve solutions?

Module 6: Lesson 4 Assignment

You will complete the Module 6 Assignment 4 in this lesson.

Remember that the questions that are not marked by the teacher provide you with the practice and feedback that you need to successfully complete this course. You should respond to all the questions and place those answers in your course folder.

Explore

Read

As you learned in Module 5, quantitative aspects of a solution include the concentration of solute and volume of solution. Can you recall a relationship that exists between a solution's concentration, its volume, and the number of moles of solute it contains? What might be similar about how you would use this relationship and how you have used other relationships, such as $n = \frac{m}{M}$ and $PV = nRT$, during your study in this module?

Read "7.4 Solution Stoichiometry" on pages 300 to 302 of your textbook. Work through "SAMPLE problem 7.4" and the "COMMUNICATION example."

Discuss

Look at the "Stoichiometry Calculations" in the margin on page 302 of your textbook. Construct a flow chart that outlines the steps involved in stoichiometry calculations. The flow chart should summarize the use of the formulas you have used to perform calculations regardless of whether the known or desired substance in a reaction is a solid, gas, or part of a solution.

Save a copy of your answers to your course folder. You may wish to share your answers with your classmates or some other people.

Self-Check

Complete "Practice" questions 1 to 3 on page 302 of your textbook.

Check your work.

Self-Check Answers

"Practice" questions 1 to 3, page 302

1.

$H_2SO_4(aq)$	+	$2\ NH_3(aq)$	→	$(NH_4)2SO_4(aq)$
$c = ?$ $V = 50.0\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}}$ $= 0.0500\ L$		$c = 2.20\ mol/L$ $V = 24.4\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}}$ $= 0.0244\ L$		

$$c = \frac{n}{V}$$

$$\begin{aligned} n_{NH_3} &= cV \\ &= (2.20\ mol/L)(0.0244\ L) \\ &= 0.053\ 68\ mol \end{aligned}$$

$$\begin{aligned} n_{H_2SO_4} &= n_{NH_3} \times \frac{1}{2} \\ &= 0.053\ 68\ mol \times \frac{1}{2} \\ &= 0.026\ 84\ mol \end{aligned}$$

$$\begin{aligned} c &= \frac{n}{V} \\ &= \frac{(0.026\ 84\ mol)}{(0.0500\ L)} \\ &= 0.537\ mol/L \end{aligned}$$

The concentration of the sulfuric acid solution is 0.537 mol/L.

2.

$3\ Ca(OH)_2(aq)$	+	$Al_2SO_4(aq)$	→	$3\ CaSO_4(s)$	+	$2\ Al(OH)_3(s)$
c = 0.0250 mol/L V = ?		$c = 0.125\ mol/L$ $V = 25.0\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}}$ $= 0.0250\ L$				

$$c = \frac{n}{V}$$

$$V = \frac{n}{c}$$

$$= \frac{(0.009\ 375\ \text{mol})}{(0.0250\ \text{mol/L})}$$

$$= 0.375\ \text{L or } 375\ \text{mL}$$

$$c = \frac{n}{V}$$

$$n_{Al_2SO_4} = cV$$

$$= (0.125\ \text{mol/L})(0.0250\ \text{L})$$

$$= 0.003\ 125\ \text{mol}$$

$$n_{Ca(OH)_2} = n_{Al_2SO_4} \times \frac{3}{1}$$

$$= 0.003\ 125\ \text{mol} \times \frac{3}{1}$$

$$= 0.009\ 375\ \text{mol}$$

The volume of the calcium hydroxide solution required is 375 mL.

3.

2 $FeCl_3(aq)$	**+**	**3 $Na_2CO_3(aq)$**	**→**	**6 $NaCl(aq)$**	**+**	**$Fe_2(CO_3)_3(s)$**
$c = 0.200\ \text{mol/L}$ $V = 75.0\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}}$ $= 0.0750\ \text{L}$		c = 0.250 mol/L V = ?				

$$C = \frac{n}{V}$$

$$n_{FeCl_3} = CV$$

$$= (0.200\ \text{mol/L})(0.0750\ \text{L})$$

$$= 0.0150\ \text{mol}$$

$$n_{Na_2CO_3} = n_{FeCl_3} \times \frac{3}{2}$$

$$= 0.0150\ \text{mol} \times \frac{3}{2}$$

$$= 0.0225\ \text{mol}$$

$$c = \frac{n}{V}$$

$$V = \frac{n}{c}$$

$$= \frac{(0.0225\ \text{mol})}{(0.250\ \text{mol/L})}$$

$$= 0.0900\ \text{L or } 90.0\ \text{mL}$$

The volume of the sodium carbonate solution required is 90.0 mL.

Module 6: Lesson 4 Assignment

Use the following lab activity to complete the first question of the Module 6 Lesson 4 Assignment.

To answer the question in an online quiz, click on Module 6 Lesson 4 Assignment in the "Quizzes".

To answer the question as an MSWord document, click **Module 6 Assignment 4**.

Lab: Analysis of Silver Nitrate

View the lab activity, **Lab Analysis of Silver Nitrate**.

Module 6: Lesson 4 Assignment

Answer the remaining questions of the Module 6 Lesson 4 Assignment.

Reflect on the Big Picture

As you have seen throughout this module, launching and operating the Space Shuttle relies on the knowledge of stoichiometry. Can you think of an example where knowledge of solution stoichiometry could be applied?

Assessment

- Module 6: Lesson 4 Assignment

Lesson Summary

In this lesson you explored the following essential question:

- How is the stoichiometric method applied to reactions that involve solutions?

You learned to perform calculations using values for concentration and volume of aqueous solutions involved in chemical reactions. Your ability to complete stoichiometric calculations for aqueous solutions can be combined with the skills you learned for solids in Lesson 2 and gases in Lesson 3.

Your study in this lesson has also allowed you to see situations where solution stoichiometry is used to finish materials, make new substances, and monitor and test solutions.

Module Glossary

gravimetric analysis: the determination of masses of substances

law of conservation of mass: a law stating that in any physical or chemical system, the initial mass of the system will be identical to the final mass of the system

mole ratio: a mathematical statement of the proportion of each substance involved in a chemical process relative to one another

net ionic equation: a type of balanced chemical equation that lists only the reacting particles

stoichiometry: a method of predicting or analyzing the amounts of the reactants and products participating in a chemical process

yield of a reaction: a measured amount of product obtained by a chemical reaction, often expressed as a percentage of maximum yield

Module Summary and Assessment

Module Summary

Stoichiometry is a method of predicting quantitative relationships within chemical systems. Knowing the amount of substances involved in a process is important when planning or analyzing a process that will involve a chemical change.

Think back to the questions you were asked at the beginning of the module:

- How do scientists and engineers use mathematics when analyzing chemical change?
- How is a balanced chemical equation used to predict amounts of species involved in a chemical process?

In this module you viewed a balanced chemical equation as a recipe for a process, where coefficients of the species involved describe the amounts of particles involved. You also used your knowledge of mathematical relationships to calculate the amount of matter participating in a process from measurable data—including mass, concentration, volume, and pressure. Finally, you verified the predictions made using the stoichiometric method by collecting and analyzing evidence obtained through laboratory investigations.

Module Assessment

The assessment in this module consists of the following:

- Assignments for Lessons 1, 2, 3, and 4