UNIT 6

Quantifying Chemical Reactions

Unit Overview

Chemists use numbers and symbols in the form of chemical formulas to represent chemical reactions. Atoms, elements, and molecules combine together in whole number ratios relative to each other. These whole number ratios are determined using physically measurable quantities, and these measurable quantities contain a huge number of atoms. This unit focuses on quantifying chemical reactions by exploring moles, atomic and molar masses, balanced chemical equations, and reaction yields. Being able to quantify chemical reactions is important in many practical applications of chemistry.

Learning Objectives and Applicable Standards

Participants will be able to:

1. Differentiate between empirical and molecular formulas and understand how they are determined experimentally.

2. Balance equations for basic chemical reactions.

3. Calculate chemical quantities from measurable quantities using mole ratios.

4. Understand how limiting reagents and incomplete reactions influences reaction yields.

Key Concepts and People

1. **Empirical Formula:** The simplest whole number ratio of elements in a compound (the empirical formula) can be determined by analyzing the products of combustion of the compound.

2. **Molecular Formula:** For many compounds, the empirical formula alone is not enough to determine the identity or structure of the compound. The molecular formula contains the actual number of each type of atom in a compound.

3. **Avogadro's Number and Atomic Mass:** The mole is the basic unit in chemical calculations. The mole represents physically measurable quantities of atoms and molecules, and thus is a huge number. The officially defined quantity of a mole is the number of carbon-12 atoms in exactly 12 grams of carbon-12.

4. **Molecular Masses and Mass Percentages:** The molecular mass of a given molecule is the sum of the atomic masses of its components. Mass percentages reveal the

fraction of mass that each type of atom contributes to the molecular mass of the overall molecule.

5. **Chemical Reactions and Equations:** Chemical equations may be written either in formulas or in chemical structures. Equations must be balanced to account for every atom on each side. A balanced equation is the basis of all reaction analysis.
6. **Stoichiometry:** The amounts of reactants and products used in a reaction can be compared in mole ratios using the coefficients of the balanced equation, and converted to measurable quantities as needed.
7. **Limiting Reagents**: Most reactions are not performed with the exact amounts of reactants necessary to consume each other completely; one or more usually will be left over. The maximum possible yield of a reaction is based on the limiting reagent.
8. **Percent Yield**: In a given chemical reaction, the percent yield reveals how much of the reactants were converted to products compared to what would be expected if 100% of the reactants were converted to 100% of the products.

Video

Stoichiometry gives us the quantitative tools to figure out the relative amounts of reactants and products in chemical reactions. Balancing the number of atoms on each side of the equation, calculating the amount of each reactant, and figuring out which reactant will run out first are all fundamental principles when designing any chemical reaction. These principles are applied when splitting water into hydrogen and oxygen for energy, manufacturing sodium iodide for radiation detectors, and producing common chemicals from renewable resources.

VIDEO CONTENT

Host Introduction
"Quantifying Chemical Reactions"

The Nocera Laboratory at Harvard University researches how to create cheap, clean, and renewable energy. Host David Song, a graduate student in the Nocera Laboratory, introduces the idea that a gallon of water could provide enough energy to power a house for an entire day.

Current Chemistry Research
"A Balanced Equation for Fuel"

The Nocera Laboratory at Harvard University has developed the "artificial leaf," a new technology that mimics photosynthesis and generates hydrogen gas from water. This system is designed to generate hydrogen for energy use by using sunlight to "split" water. The amount of hydrogen produced in the process can be calculated based on the balanced equation for the chemical reaction.

Laboratory Demonstration
"Making Water 2:1"
The combination of hydrogen gas and oxygen gas to form water requires twice as much hydrogen as oxygen, but what happens if you mix equal parts of the two gases? Dr. Wolfgang Rueckner, Manager of Lecture Science Demonstration Services at Harvard University, shows what happens.

Host Science Explanation
"Calculating Molar Mass"
In addition to a balanced equation, chemists need a way to deal with extremely large quantities of atoms and molecules in order to calculate chemical quantities. David Song introduces the concept of the mole and uses the reaction that forms sodium iodide, a compound used in many radiation detectors, to show how to calculate molar mass.

Host Introduction to Science Explanation
"Limiting Reagent"
Usually, one reagent in a reaction is consumed completely before the remaining reagents are exhausted. This reagent is referred to as the limiting reagent. Dave Song demonstrates this concept by baking cookies in which the chocolate chips serve as the limiting reagent.

Laboratory Demonstration
"Finding the Limiting Reagent"
Methanol vapor will ignite on a wire of catalytic metal when allowed to react in the presence of oxygen. Harvard University Lecture Demonstrator Daniel Rosenberg shows that when the methanol vapor is reacted in an open container, the reaction proceeds continuously. When the container is closed, however, the reaction ultimately stops due to a lack of oxygen before all the methanol is consumed.

Real-World Application
"Sustainable Chemistry"
Myriant Corporation, a global renewable chemicals company located in the Boston area, uses renewable resources to make chemical products that would ordinarily be made from the non-renewable resource, petroleum. Researchers at Myriant show that in order to produce environmentally friendly chemical products, it is important to understand how to balance chemical equations, how to calculate molecular masses, and how to work to your advantage with limiting reagents.

Unit Text

Content Overview
This unit introduces the basics of stoichiometry: how chemists are able to determine molecular quantities of reactants and products in chemical reactions. First, the text explains that atoms combine in whole number ratios to form molecules, which are represented by empirical and molecular formulas. In order to deal with the large quantities of atoms involved in a given

chemical reaction, chemists use a unit of measurement called the "mole" to represent amounts of a chemical substance. The text then explains Avogadro's number, atomic masses, and molar masses. This information leads to a discussion of chemical reactions and how they are represented by balanced chemical equations, followed by an explanation of stoichiometry. The unit ends by discussing limiting reagents and percent yields of reactions.

Sidebar Content

1. **Stinking Up the House:** Oliver Sacks recounts a time when he produced a large amount of hydrogen sulfide in his house.
2. **Berzelius and the Birth of Formulas:** Jons Jacob Berzelius is the man behind our current system of writing chemical formulas.
3. **Isomers and Lewis Structures:** Butanol and diethyl ether are isomers, which means they have the same molecular formula, but have different structures, and thus very different properties. Lewis structures help represent the differences in the isomers' structures.
4. **How Many Items Are in a Mole?** This sidebar shows examples of how big a number one mole really is.
5. **Avogadro's Law Example:** Avogadro's 1811 theory predicted that a given volume of any gas has the same number of particles.
6. **Converting Moles to Atomic Mass:** There are two major ways to use atomic masses to determine how much of an element to use in a given chemical reaction.
7. **Equilibrium Reactions:** Even a small amount of carbon dioxide dissolved in water (even a huge body of water!) can make the solution acidic by the formation of carbonic acid.

Interactives

Historical Timeline of Chemistry
This interactive illustrates how different discoveries build upon, disprove, or reinforce previous theories. This not only reinforces basic chemistry concepts, but also emphasizes the nature of science. Scientists mentioned in this unit are listed on the timeline.

Chemistry of Running Interactive
Equilibrium, particularly the equilibrium of hemoglobin and oxyhemoglobin, heavily influences the physiology of exercise. Although many biochemical reactions have very simple stoichiometry, the actual amounts of the various reactants and products can heavily influence the body's ability to perform strenuous activity.

Control a Haber-Bosch Ammonia Plant Interactive

The Haber process is based on an equilibrium process. Reaction yields are dependent not only upon the reaction conditions overall, but also fundamentally upon the stoichiometry of the reaction. For example, for the production of two molecules of ammonia, one molecule of nitrogen and three molecules of hydrogen are required. Please note that a lesson plan and student worksheet are available online with this interactive.

During the Session

Before Facilitating this Unit

Few topics are as critical to chemistry as the mole concept, but few topics are more intimidating to students. Often, so much emphasis is placed on the sheer size of Avogadro's number that students don't realize how rarely they will actually need to use it directly. They invoke it far too frequently in their calculations. The focus in this unit is on chemical proportions and on expressing chemical quantities in terms that we can easily describe.

Tips and Suggestions

1. Although Avogadro's number is intimidatingly large, **it is actually not needed as frequently in calculations as students often think**. Most calculations in chemistry use mole proportions with respect to measurable quantities such as masses. While the scale of the number is important, and the actual value is used under some circumstances, the focus of sample calculations should be slanted more toward mole ratios in compounds and equations, and the relationship of those quantities to masses, volumes, and other readily measured parameters. Encourage students to first set up multi-step calculations in terms of the kinds of quantities they are attempting to convert, and then to include the numerical values once they have developed a strategy.

2. **The empirical formula for a compound is not sufficient to determine the identity of a molecular compound**, but even when the molecular formula is known, many molecular formulas can be drawn with different isomeric structures. A discussion of how much information can be obtained from a given type of experiment is useful not only for introduction of the different types of formulas but also for introduction to the link between experiment design, execution, and conclusions.

3. **The determination of a limiting reagent may be done in various ways**, but the simplest way is to calculate how much of the desired product can be formed in a reaction using each of the known quantities of reactant. Even if one reagent is obviously limiting (e.g., a reaction with oxygen is performed in open air), it is useful to emphasize the limiting reagent so that students acquire the habit of looking for it before beginning calculations.

Starting the Session: Checking Prior Thinking

You might assign students a short writing assignment based on the following questions, and then spend some time discussing prior thinking. This will help elicit prior thinking and misconceptions.

1. How many molecules of water do you think are in a given glass of water?
2. What are some ways chemists measure how much of a given compound they have?
3. What does a chemical equation represent?
4. What does this chemical equation show: $2H_2 + O_2 \rightarrow 2H_2O$? Draw a visual representation of this.
5. Is it important for chemical equations to be balanced? Why or why not?
6. Which is heavier: one atom of iron or one atom of gold? If you have 10g each of iron and gold, which one will contain more atoms?
7. Suppose you had 50 atoms of iron and 50 molecules of water. Which sample would weigh more? Which sample would contain more atoms total?
8. If you have 20 atoms of hydrogen and 10 atoms of oxygen, how many molecules of water can you make? What if you have 20 atoms of hydrogen and 20 atoms of oxygen? What if you have 40 atoms of hydrogen and 10 atoms of oxygen?

Before Watching the Video

Students should be given the following questions to consider while watching:

1. What is the balanced chemical equation for splitting water?
2. What happens when two parts hydrogen and two parts oxygen combine?
3. What does a mole represent?
4. How do you use the periodic table to determine the molar mass of sodium iodide (NaI)?
5. What is a limiting reagent?
6. In the reaction of methanol vapor on the catalytic wire what reaction is happening? How do you know methanol is the limiting reagent?

Watch the Video

After Watching the Video

Use these additional questions as follow-up, either as a group discussion or as short writing assignments.

1. Why is having a balanced reaction important in chemical calculations?

2. What is the ideal way to minimize waste when performing chemical reactions? If it is not possible to eliminate waste entirely, what factors might influence the strategy to minimize it?

3. Are masses like grams and kilograms always the easiest way to measure chemical quantities? How else might you need to measure amounts of reactants and products in order to perform chemical calculations?

4. Many reactions, especially in industry, are monitored for their progress over time. Is it necessary (or sensible) to do this by measuring the amount of reactant still present? Why or why not?

Group Learning Activities

Building Isomers

Objective

For this activity, students will use models to explore different isomeric forms of chemical formulas. For example, C_2H_6O may be the formula for either ethanol or dimethyl ether, and the models of the two structures may be switched simply by the interchange of the –H and the $–CH_3$.

List of Materials

- Chemical modeling kits
- If chemical modeling kits aren't available, marshmallows and toothpicks can be used to model molecules. For example, large marshmallows can represent carbon atoms, small marshmallows can be hydrogen atoms, and different colored marshmallows or other soft candies can be other atoms. Although marshmallows do not hold bond angles particularly firmly, the models should give students a reference from which to draw the Lewis structures of different isomers, or vice versa.

Set Up

List molecular formulas on the board. Suitable molecular formulas include, but are not limited to: C_2H_6O, C_6H_{14}, $C_4H_{11}N$, $C_6H_{12}O$, and $C_6H_{13}Br$.

Procedure

1. If using molecular modeling kits, connect atoms to build as many different configurations of molecules represented by the molecular formulas on the board, keeping in mind the limit of how many bonds each atom can make.

2. If using marshmallows and toothpicks, gather enough large marshmallows (for carbon atoms), small marshmallows (for hydrogen atoms), and colored marshmallows or other candies (for atoms other than C and H) to account for every atom in one of the

molecular formulas listed on the board. Using toothpicks to connect the marshmallows, build as many different configurations as possible of the different atoms, keeping in mind the limit of how many bonds each atom can make.

3. From each model, draw a Lewis structure of the compound.

Discussion
The following questions can guide a discussion and help students reflect on this exercise.

1. Which molecular formula gave the most possible isomers? Which one gave the fewest?

2. Did you have to dismantle the models completely in order to build new isomers of the same compound? Explain.

3. Why do you think different isomers have different chemical properties?

Hazards
There are no hazards associated with this exercise; however, if the exercise is performed on a laboratory bench, the candies should not be eaten afterwards.

Disposal
Marshmallows and toothpicks may be disposed in the trash.

Building Reagents and Products with Toys
(May also be done as a demo)
Objective
The variations of this exercise all address different principles of chemical proportions: specific numbers of atoms are necessary to form specific compounds, but in some cases the same numbers and types of atoms may be able to combine in different ways. In all cases, simple models (not necessarily molecular models) are built from commercially available construction toys.

List of Materials
- LEGO® Bricks, blocks, TINKERTOYS® or other construction toys; each group should be given enough pieces to construct a few model structures with some extra pieces left over.

Set Up for Variation I: Limiting Reagent Part 1
Build simple models/structures from the building materials for the students to copy. For example, you could build a structure with one red brick and two small yellow bricks attached to one side, etc.

Procedure for Variation I: Limiting Reagent Part 1
1. Using the pieces available in your set, build as many replicas as possible of the model displayed at the instructor's desk.

2. Once you have built as many of the models as your available pieces will allow, make a note of which piece(s) remain and which were completely consumed in the building of the target models.

3. Record what the limiting reagents were for each model.

Set Up For Variations II and III
As indicated for variation I, build models/structures for the students to copy, but instead of providing individual building pieces to the students to build their models, provide smaller, simple models/structures as starting materials.

Variation II: Limiting Reagents Part 2
1. Students will be shown simple models of "reactants" and "products" at the instructor's desk, and given several of each model of the different "reactant" constructions.

2. Using only the pieces available in your set, build as many sets of the models at the instructor's desk as possible. Dismantle models from your set only if you will be able to construct "product" models and not leave behind portions of the original "reactant" models.

3. Once you have built as many of the "product models" as your available "reactant" models will allow, make a note of which original model(s) remain and which were completely consumed. Which were the limiting reagents?

Variation III: Reactants to Products
1. You will be given two or more pre-constructed simple models.

2. Rearrange the pieces from your given models into as many different combinations of new models as possible, with no individual pieces left over.

3. Make a note of each combination you come up with before dismantling the models to generate new combinations.

Discussion
These questions can help guide students' thinking during and after the demo:

1. In Variation I: Limiting Reagent Part 1, what represents the limiting reagent? How do you know?

2. The building of models from individual blocks is analogous to the formation of compounds from individual atoms, while the building of models from components of other models is analogous to the formation of compounds from existing compounds. Which is more common for actual chemical reactions? Explain.

3. In variation III, you could have created "products" from the original models you were

given by dismantling all the models completely into individual blocks. How is this similar to decomposition of molecules into their component elements? How is it different?

Hazards
There are no hazards associated with this activity.

Disposal
The materials for this exercise are reusable.

In-Class Chemical Demonstrations

When Oxygen Is Limiting

Objective
In this demonstration, the burn times for a candle under a flask of air and under a flask of pure oxygen are compared and used to estimate the percentage of oxygen in the atmosphere.

Materials Needed
- Two short taper candles, at least one of a diameter and height to be able to fit under an inverted 250mL Erlenmeyer flask
- Two 250mL Erlenmeyer flasks
- Oxygen tank or lecture bottle, with rubber tubing to bubble oxygen into underwater flask
- Basin to fill with water
- Stopper for one of the flasks
- Stopwatch
- Matches or lighter

Procedure
1. Fill the basin with water and submerge one of the two flasks to fill it.

2. Open the regulator on the oxygen tank and bubble oxygen into the submerged flask through the hose.

3. Close the flask with the stopper, and then remove the flask from the basin.

4. Dry the outside of the oxygen-filled flask.

5. Secure the taper candles to stand up (either use a candle tray or soften the wax at the base to stick to the table). Light both candles.

6. Cover one of the candles with the NON-stoppered Erlenmeyer flask (*i.e.*, the one not filled with pure oxygen). Have students measure and record the time from the covering of the candle to the extinguishing of the flame.

7. Once the flame is extinguished and data are recorded, light the candle again, but this time cover the candle with the flask containing pure oxygen. Again, have students measure and record the time required to extinguish the flame under the flask.

8. Point out to the students that the second candle (not put under the flasks) has been burning the entire time.

Discussion

These questions can help guide students thinking during and after the demo:

1. Why did the candle on the desk stay lit while the candles put under flasks went out?

2. Calculate the ratio between extinguishing times for the candle under the flask filled with air to the candle under the flask filled with pure oxygen. How close does this ratio come to 21% (the approximate percentage of the atmosphere occupied by oxygen)?

3. If the flasks used to cover the candles had been twice as large, what do you think the lengths of time to extinguish the candles would be? Explain.

Hazards

Gas cylinders must be secured appropriately before use. Keep the gas cylinder away from the matches/lighters and lighted candles. Be careful not to burn yourself when lighting candles. Oxygen is highly flammable.

Disposal

Candles are reusable. Wet matches before disposing in the trash.

Going Deeper (In-class Discussion or Reflection)

Instructors should allow up to 30 minutes for discussion at the end of the session, or students can use the time to reflect on one or more of these questions in journals.

1. What happens to the mass of material inside an open container if a reaction releases gas-phase products? What about in a closed container? What happens to the pressure inside a closed container if a reaction releases gas-phase products? What about in an open container?

2. A compound containing only C and H is burned for combustion analysis. If the reaction produces equal moles of carbon dioxide and of water, what is the empirical formula of the compound? If the compound's formula weight is 84 g/mol, what is the molecular formula of the compound?

3. You are working in a lab, and need to synthesize 10 g of a compound. The procedure says that the reaction you are using gives only a 70% yield. How much of the compound do you need to plan to make in order to have an actual yield of 10 g?

4. You are doing a multi-step procedure in a lab. Step 1 has a 70% yield, step 2 has a 20% yield, and step 3 has a 95% yield. What will the total percent yield for the process be?
5. Suppose you worked in a lab and were studying the synthesis described above. If you were trying to develop a way to improve the overall yield of the synthesis, which step would you first attempt to improve, and why?
6. Suppose you are setting up a reaction and need to determine which reagent should be limiting. What factors should you consider in choosing the limiting reagent?
7. In many chemical reactions, more than one product is produced. If only one of the products is desired, what are some of the ways that the product might be isolated?

Before the Next Unit

Learners should read the Unit 6 text if they haven't already done so. They may wish to read one or more of the reading assignments from the list below, or, if you choose to have them use the course materials outside of class, they can watch the Unit 7 video and/or read the Unit 7 text as an assignment before the next session.

References and Additional Resources

Thompson, Stephen. *Chemtrek: Small-Scale Experiments for General Chemistry*. Vol 2. Boston: Allyn and Bacon, 1990.

Cacciatore, Kristen L. and Hannah Sevian. "Teaching Lab Report Writing through Inquiry: A Green Chemistry Stoichiometry Experiment for General Chemistry." *Journal of Chemical Education*. 83: 7 (2006). 1039-1041. Accessed August 9, 2013. http://pubs.rsc.org/en/content/articlelanding/2007/rp/b6rp90017h/unauth#!divAbstract

Fach, Martin, Tanja de Boer and Ilka Parchmann. "Results of an Interview Study as Basis for the Development of Stepped Supporting Tools for Stoichiometric Problems." *Chemistry Education Research and Practice*. 8:1 (2007) 13-31. Accessed August 9, 2013. http://www.rsc.org/images/Fach%20paper%20final_tcm18-76278.pdf

For Professional Development

In addition to watching the videos, reading the text, and going through the activities listed in the course guide, participants taking this course for professional development should read the following papers and answer the corresponding reflection questions. Participants should then complete the accompanying professional development assignments.

Further Reading & Reflection Questions

DeMeo, Stephen. "Revisiting Molar Mass, Atomic Mass, and Mass Number: Organizing, Integrating, and Sequencing Fundamental Chemical Concepts." *Journal of Chemical Education*. 83:4 (2006). 617-621. Accessed August 9, 2013. http://www.bruderchemistry.com/core/content/files/Revisiting%20Mass%20article%201.pdf

1. Did this paper convince you that it makes the most sense to teach molar mass first, then atomic mass, then mass number? Why or why not? What do you think are some benefits and challenges to teaching the content in this order?

2. Do you think the questions presented in "The Creation of New Problems" could assess a student's conceptual understanding of molar mass, atomic mass, and mass number? Why or Why not?

3. Did this paper help clarify any confusion you may have had with these terms? If so, how?

Arasasingham, Ramesh D., Mare Taagepera, Frank Potter, Stacy Lonjers. "Using Knowledge Space Theory to Assess Student Understanding of Stoichiometry." *Journal of Chemical Education*. 81: 10 (2004). Accessed August 9, 2013. http://pubs.rsc.org/en/Content/ArticleLanding/2011/RP/c1rp90039k#!divAbstract (article available via free login)

1. Were you surprised by the results of this study? Why or why not? What was most surprising and why? What was least surprising and why?

2. Have you encountered students that held misconceptions similar to the ones revealed in the study? What are some ways you could address these types of misconceptions?

3. Do you agree with the authors' suggestions for how to improve the teaching of these concepts? How might you apply some of their suggestions to how you teach these materials?

Sanger, Michael J. "Evaluating Students' Conceptual Understanding of Balanced Equations and Stoichiometric Ratios Using a Particulate Drawing." *Journal of Chemical Education*, v82 n1 p131 Jan 2005. Accessed October 24, 2014. http://pubs.acs.org/doi/abs/10.1021/ed082p131 (available via ACS) and http://eric.ed.gov/?id=EJ726143

1. Were you surprised by the results of this study? Why or why not? What was most surprising to you and why?

2. How have you assessed students' knowledge of stoichiometric principles in the past? In teaching stoichiometry, have you encountered students who are able to solve stoichiometric calculations, but who lacked conceptual understanding of stoichiometry? How have you addressed this?

3. Does this study influence how you will teach or assess stoichiometric principles? If so, how? If not, why not?

Professional Development Assignments

1. After reading the papers above and reflecting on the questions presented develop a lesson plan designed to teach material presented in this unit.

2. Using a group activity or classroom demonstration presented in this course guide, show how you would implement it into your classroom. Where would it fit into your curriculum or standards? Would you change the demonstration or activity in any way? How would you assess student learning?