12.2 Chemical Calculations

Connecting to Your World
Air bags inflate almost instantaneously upon a car’s impact. The effectiveness of air bags is based on the rapid conversion of a small mass of sodium azide into a large volume of gas. The gas fills an air bag, preventing the driver from hitting the steering wheel or dashboard. The entire reaction occurs in less than a second. In this section you will learn how to use a balanced chemical equation to calculate the amount of product formed in a chemical reaction.

Writing and Using Mole Ratios
As you just learned, a balanced chemical equation provides a great deal of quantitative information. It relates particles (atoms, molecules, formula units), moles of substances, and masses. A balanced chemical equation also is essential for all calculations involving amounts of reactants and products. For example, suppose you know the number of moles of one substance. The balanced chemical equation allows you to determine the number of moles of all other substances in the reaction.

Look again at the balanced equation for the production of ammonia from nitrogen and hydrogen:

\[ \text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g) \]

The most important interpretation of this equation is that 1 mol of nitrogen reacts with 3 mol of hydrogen to form 2 mol of ammonia. Based on this interpretation, you can write ratios that relate moles of reactants to moles of product. A mole ratio is a conversion factor derived from the coefficients of a balanced chemical equation interpreted in terms of moles. In chemical calculations, mole ratios are used to convert between moles of reactant and moles of product, between moles of reactant, or between moles of products. Three mole ratios derived from the balanced equation above are:

\[ \frac{1 \text{ mol N}_2}{3 \text{ mol H}_2} \quad \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} \quad \frac{3 \text{ mol H}_2}{2 \text{ mol NH}_3} \]

Mole-Mole Calculations In the mole ratio below, \( W \) is the unknown quantity. The values of \( a \) and \( b \) are the coefficients from the balanced equation. Thus a general solution for a mole-mole problem, such as Sample Problem 12.2, is given by

\[
x \text{ mol } G \times \frac{b \text{ mol } W}{a \text{ mol } G} = \frac{xb}{a} \text{ mol } W
\]

Given Mole ratio Calculated

Guide for Reading

Key Concepts
- How are mole ratios used in chemical calculations?
- What is the general procedure for solving a stoichiometric problem?

Vocabulary
mole ratio

Reading Strategy
Relating Text and Visuals As you read, look closely at Figure 12.4. Explain how this illustration helps you understand the relationship between known and unknown quantities in a stoichiometric problem.

Figure 12.4 Manufacturing plants produce ammonia by combining nitrogen with hydrogen. Ammonia is used in cleaning products, fertilizers, and in the manufacture of other chemicals.

Guide for Reading

Build Vocabulary

Paraphrase Have students work with a partner to define the term mole ratio in their own words. They may do so by reading this section and by using what they have already learned about balanced chemical equations. Have student pairs read their definitions to the class.

Reading Strategy
Identify Main Ideas/Details As you read the material under the heading Mass-Mass Calculations, identify and list the main ideas presented by the text.

INSTRUCT

Connecting to Your World
Have students study the photograph and read the text that opens the section. Write the equation for the decomposition of sodium azide with heat as one of the products \( 2\text{NaN}_3(s) \rightarrow 2\text{Na}(s) + 3\text{N}_2(g) + \text{heat} \). Ask, How can stoichiometry be used to calculate the volume of a gas produced in this reaction? (The number of moles (and volume) of nitrogen gas formed by this reaction depends on the number of moles of sodium azide that decompose.)

Section Resources
Print
- Guided Reading and Study Workbook, Section 12.2
- Core Teaching Resources, Section 12.2 Review
- Transparencies, T126–T132
- Laboratory Manual, Lab 19

Technology
- Interactive Textbook with ChemASAP, Simulation 13, Problem-Solving 12.12, 12.13, 12.15, 12.19, Assessment 12.2

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Section 12.2 (continued)

Writing and Using Mole Ratios

Using Visuals

Figure 12.4 Have students consider the sizes of the containers shown relative to the sizes of containers used in a classroom laboratory. Have them imagine they are managing the manufacturing facility pictured. Ask, What factors would they need to consider to meet demands for ammonia? (Acceptable answers include the number of customers, the number of cylinders per customer, the amount of ammonia per cylinder, and the amount of H₂ and N₂ needed to produce that quantity of NH₃.)

Sample Problem 12.2

Answers

11. a. 4 mol Al₂O₃ 3 mol O₂ 4 mol Al
   2 mol Al₂O₃ 3 mol O₂ 2 mol Al₂O₃
   4 mol Al
   b. 7.4 mol

12. a. 11.1 mol
   b. 0.52 mol

Practice Problems Plus

Chapter 12 Assessment problem 38 is related to Sample Problem 12.2.

Discuss

Point out that heat is produced in the decomposition of sodium azide, used in air safety bags. Explain the heat produced by a reaction can also be measured and related to the amount of reactant(s) consumed. In this case, the amount of heat produced depends on the mass of sodium azide (the reactant) that decomposes. It is possible to relate grams of reactant to moles of reactant to heat produced.

Math Handbook

For a math refresher and practice, direct students to dimensional analysis, page R66.

Differentiated Instruction

Less Proficient Readers

Encourage students to find a method of problem solving that capitalizes on their strengths. For example, a visual learner might draw pictures of the reactants and products, instead of just writing the symbols. A kinesthetic learner may prefer to manipulate molecular models of the reactants and products.

360 Chapter 12
**SAMPLE PROBLEM 12.3**

**Calculating the Mass of a Product**

Calculate the number of grams of NH₃ produced by the reaction of 5.40 g of hydrogen with an excess of nitrogen. The balanced equation is

\[
N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)
\]

1. **Analyze** List the knowns and the unknown.

- **Knowns**
  - Mass of hydrogen = 5.40 g H₂
  - 3 mol H₂ = 2 mol NH₃ (from balanced equation)
  - 1 mol H₂ = 2.0 g H₂ (molar mass)
  - 1 mol NH₃ = 17.0 g NH₃ (molar mass)

- **Unknown**
  - Mass of ammonia = ? g NH₃

The mass in grams of hydrogen will be used to find the mass in grams of ammonia:

\[
g\text{H}_2 \rightarrow g\text{NH}_3
\]

The following steps are necessary to determine the mass of ammonia:

\[
g\text{H}_2 \rightarrow \text{mol H}_2 \rightarrow \text{mol NH}_3 \rightarrow g\text{NH}_3
\]

The coefficients of the balanced equation show that 3 mol H₂ reacts with 1 mol N₂ to produce 2 mol NH₃. The mole ratio relating mol NH₃ to mol H₂ is 2 mol NH₃/3 mol H₂.

2. **Calculate** Solve for the unknown.

This following series of calculations can be combined:

\[
5.40 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.0 \text{ g H}_2} \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} \times \frac{17.0 \text{ g NH}_3}{1 \text{ mol NH}_3} = 31 \text{ g NH}_3
\]

3. **Evaluate** Does the result make sense?

Because there are three conversion factors involved in this solution, it is more difficult to estimate an answer. However, because the molar mass of NH₃ is substantially greater than the molar mass of H₂, the answer should have a larger mass than the given mass. The answer should have two significant figures.

**Practice Problems**

13. Acetylene gas (C₂H₂) is produced by adding water to calcium carbide (CaC₂). C₃H₄ (g) + H₂O(l) → C₂H₂(g) + Ca(OH)₂(aq)

How many grams of acetylene are produced by adding water to 5.00 g CaC₂?

14. Using the same equation, determine how many moles of CaC₂ are needed to react completely with 49.0 g H₂O.

**Significant Figures**

The significant figures in a measurement are all the digits known with certainty plus one estimated digit. The number of significant figures in the measurements used in a calculation determines how you round the answer.

When multiplying and dividing measurements, the rounded answer can have no more significant figures than the least number of significant figures in any measurement in the calculation.

The product of 3.6 m × 2.48 m = 8.928 m² is rounded to 8.9 m² (2 significant figures).

When adding and subtracting measurements, the answer can have no more decimal places than the least number of decimal places in any measurement in the problem. The difference of 8.78 cm – 2.2 cm = 6.58 cm is rounded to 6.6 cm (one decimal place).

When determining which digits are significant, ask, How many significant figures are in each of the following condensed rules for determining which digits are significant?

1. Nonzero digits are significant.
2. A zero is significant only if it is at the right end of a number and after a decimal point, or between digits that are significant.
3. If a quantity is exact, it has an unlimited number of significant digits.
4. All digits of a quantity written in scientific notation are significant.

**Answers**

13. 2.03 g C₂H₂
14. 1.36 mol CaC₂

**Practice Problems Plus**

Rust (Fe₃O₄) is produced when iron (Fe) reacts with oxygen (O₂).

4Fe(s) + 3O₂(g) → 2Fe₃O₄(s)

How many grams of Fe₃O₄ are produced when 12.0 g of iron rusts?

(17.2 g)
TEACHER Demo

Interpreting a Chemical Equation

**Purpose** Students interpret a balanced equation in terms of moles and mass.

**Materials** Prior to the demonstration prepare 0.1 M solutions of potassium iodide and lead(II) nitrate. Measure 50.0 mL of Pb(NO₃)₂ and 150 mL of KI into separate 250-mL beakers.

**Safety** Wear safety glasses and apron.

**Procedure** Tell students that you are going to mix 0.005 moles of lead(II) nitrate with excess potassium iodide. Have student observe as you combine both solutions in the 250-mL beaker. Have students write a balanced chemical equation for the observed reaction.

\[ 2\text{KI(aq)} + \text{Pb(NO₃)₂(aq)} \rightarrow 2\text{KNO₃(aq)} + \text{PbI₂(s)} \]

Have students predict the number of moles of product produced. (0.005 moles PbI₂ assuming the reaction was complete) Note that, in an actual reaction, the amounts of reactants often are not present in the mole ratios predicted by the coefficients in a balanced equation. Explain the importance of the mole ratios in an equation for calculating relative quantities. Ask, What is the mass of lead(II) nitrate reacted and the mass of lead(II) iodide produced? (1.66 g Pb(NO₃)₂ and 2.30 g PbI₂)

**Expected Outcome** A bright yellow precipitate will form.

**Discuss**

Students sometimes try to do mass-mass conversions by incorrectly using the mole ratio as a mass ratio. (That is, they use grams instead of moles as the units in the mole ratio, and then skip the mass-mole conversion step.) Stress that because the number of grams in one mole of a substance varies with its molar mass, a mass-mole conversion is a necessary intermediate step in mass-mass stoichiometric problems.

If the law of conservation of mass is true, how is it possible to make 31 g NH₃ from only 5.40 g H₂? Looking back at the equation for the reaction, you will see that hydrogen is not the only reactant. Another reactant, nitrogen, is also involved. If you were to calculate the number of grams of nitrogen needed to produce 31 g NH₃ and then compare the total masses of reactants and products, you would have an answer to this question. Go ahead and try it!

Mass-mass problems are solved in basically the same way as mole-mole problems. Figure 12.7 reviews the steps for the mass-mass conversion of any given mass \((G)\) and any wanted mass \((W)\).

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**Figure 12.6** In this Hubble Space Telescope image, clouds of condensed ammonia are visible covering the surface of Saturn.

**Figure 12.7** This general solution diagram indicates the steps necessary to solve a mass-mass stoichiometry problem: convert mass to moles, use the mole ratio, and then convert moles to mass. **Inferring** Is the given always a reactant?

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**Facts and Figures**

### Atmospheric Ammonia

Ammonia is found in trace amounts in the atmospheres of three Jovian planets—Jupiter, Saturn, and Uranus. In Jupiter’s atmosphere, the clouds of ammonia consist of frozen ammonia droplets changing to liquid ammonia droplets nearer the planet’s surface. Because of colder temperatures, the ammonia clouds in the atmosphere of Saturn and Uranus consist of frozen ammonia droplets.
Other Stoichiometric Calculations

As you already know, you can obtain mole ratios from a balanced chemical equation. From the mole ratios, you can calculate any measurement unit that is related to the mole. The given quantity can be expressed in numbers of representative particles, units of mass, or volumes of gases at STP. The problems can include mass-volume, particle-mass and volume-volume calculations. For example, you can use stoichiometry to relate volumes of reactants and products in the reaction shown in Figure 12.8. In a typical stoichiometric problem, the given quantity is first converted to moles. Then the mole ratio from the balanced equation is used to calculate the number of moles of the wanted substance. Finally, the moles are converted to any other unit of measurement related to the unit mole, as the problem requires.

Thus far, you have learned how to use the relationship between moles and mass (1 mol = molar mass) in solving mass-mass, mass-mole, and mole-mass stoichiometric problems. The mole-mass relationship gives you two conversion factors.

\[
\frac{1 \text{ mol}}{\text{molar mass}} \quad \text{and} \quad \frac{\text{molar mass}}{1 \text{ mol}}
\]

Recall from Chapter 10 that the mole can be related to other quantities as well. For example, 1 mol = 6.02 × 10^23 representative particles, and 1 mol of a gas = 22.4 L at STP. These two relationships provide four more conversion factors that you can use in stoichiometric calculations.

\[
\frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ particles}} \quad \text{and} \quad \frac{6.02 \times 10^{23} \text{ particles}}{1 \text{ mol}}
\]

\[
\frac{1 \text{ mol}}{22.4 \text{ L}} \quad \text{and} \quad \frac{22.4 \text{ L}}{1 \text{ mol}}
\]

Figure 12.8 summarizes the steps for a typical stoichiometric problem. Notice that the units of the given quantity will not necessarily be the same as the units of the wanted quantity. For example, given the mass of G, you might be asked to calculate the volume of W at STP.

**Check Point** What conversion factors can you write based on the mole-mass and mole-volume relationships?
Section 12.2 (continued)

Discuss

Construct a diagram on the board or overhead projector showing the relationships that are useful for solving stoichiometry problems. One simple model reaction is \(A \rightarrow B\). Use double-headed arrows to connect the terms: 
- Particles of \(A\)
- Moles of \(A\)
- Grams of \(A\)
- Moles of \(B\)
- Particles of \(B\)
- Grams of \(B\)
Above the appropriate arrows, write:
- Avogadro’s number
- Coefficients
- Molar mass

Explain that the only “transitions” allowed are between quantities connected by arrows. Point out that the required conversion factor to make a “transition” is written above each arrow. Have students refer to the diagram when working practice problems.

Sample Problem 12.4

Answers
15. \(4.82 \times 10^{22}\) molecules \(O_2\)
16. 11.5 g \(NO_2\)

Practice Problems Plus

Hydrogen gas can be made by reacting methane (\(CH_4\)) with high-temperature steam:

\[CH_4(g) + H_2O(g) \rightarrow CO(g) + 3H_2(g)\]

How many hydrogen molecules are produced when 158 g of methane reacts with steam? \((1.78 \times 10^{25}\) hydrogen molecules\)

SAMPLE PROBLEM 12.4

Calculating Molecules of a Product

How many molecules of oxygen are produced when 29.2 g of water is decomposed by electrolysis according to this balanced equation?

\[2H_2O(l) \xrightarrow{electricity} 2H_2(g) + O_2(g)\]

1. Analyze
   - List the knowns and the unknown.
   - Knowns
     - mass of water = 29.2 g \(H_2O\)
     - 2 mol \(H_2O\) = 1 mol \(O_2\) (from balanced equation)
     - 1 mol \(H_2O\) = 18.0 g \(H_2O\) (molar mass)
     - 1 mol \(O_2\) = 6.02 \times 10^{23} molecules \(O_2\)
   - Unknown
     - molecules of oxygen = \(?\) molecules \(O_2\)

   The following calculations need to be done:
   
2. Calculate
   - Solve for the unknown.

   \[
   \begin{align*}
   \text{Given quantity} & \quad \text{Change to moles} & \quad \text{Mole ratio} & \quad \text{Change to molecules} \\
   29.2 \text{ g } H_2O & \times \frac{1 \text{ mol } H_2O}{18.0 \text{ g } H_2O} & \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O} & \times \frac{6.02 \times 10^{23} \text{ molecules } O_2}{1 \text{ mol } O_2} \\
   & & & = 4.88 \times 10^{23} \text{ molecules } O_2
   \end{align*}
   \]

3. Evaluate
   - Does the result make sense?
   The given mass of water should produce a little less than 1 mol of oxygen, or a little less than Avogadro’s number of molecules. The answer should have three significant figures.

Practice Problems

15. How many molecules of oxygen are produced by the decomposition of 6.54 g of potassium chlorate (\(KClO_3\))? \(2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)\)
16. The last step in the production of nitric acid is the reaction of nitrogen dioxide with water. \(3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO(g)\)

The coefficients in a chemical equation indicate the relative number of particles and the relative number of moles of reactants and products. For a reaction involving gaseous reactants or products, the coefficients also indicate relative amounts of each gas. As a result, you can use volume ratios in the same way you have used mole ratios.

Differentiated Instruction

Gifted and Talented

Have students use the calculations in the sample problems as general algorithms to write computer programs that solve stoichiometric problems. Have students demonstrate and explain their programs to interested students.
SAMPLE PROBLEM 12.5

Volume-Volume Stoichiometric Calculations

Nitrogen monoxide and oxygen gas combine to form the brown gas nitrogen dioxide, which contributes to photochemical smog. How many liters of nitrogen dioxide are produced when 34 L of oxygen reacts with an excess of nitrogen monoxide? Assume conditions of STP.

\[ 2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g) \]

1. Analyze List the knowns and the unknown.
   - **Knowns**
     - volume of oxygen = 34 L \( \text{O}_2 \)
     - \( 2 \text{ mol NO}_2 / 1 \text{ mol O}_2 \) (mole ratio from balanced equation)
     - \( 1 \text{ mol O}_2 = 22.4 \text{ L} \) (at STP)
     - \( 1 \text{ mol NO}_2 = 22.4 \text{ L NO}_2 \) (at STP)
   - **Unknown**
     - volume of nitrogen dioxide = ? L \( \text{NO}_2 \)

2. Calculate Solve for the unknown.

\[
\frac{34 \text{ L O}_2 \times 1 \text{ mol O}_2}{22.4 \text{ L O}_2} \times \frac{2 \text{ mol NO}_2}{1 \text{ mol O}_2} \times \frac{22.4 \text{ L NO}_2}{1 \text{ mol NO}_2} = 68 \text{ L NO}_2
\]

3. Evaluate Does the result make sense?

Because \( 2 \text{ mol NO}_2 \) is produced for each \( 1 \text{ mol O}_2 \) that reacts, the volume of \( \text{NO}_2 \) should be twice the given volume of \( \text{O}_2 \). The answer should have two significant figures.

Practice Problems

17. The equation for the combustion of carbon monoxide is \( 2\text{CO}(g) + \text{O}_2(g) \rightarrow 2\text{CO}_2(g) \).
   How many liters of oxygen are required to burn 3.86 L of carbon monoxide?

18. Phosphorus and hydrogen can be combined to form phosphine (PH\(_3\)).
   \( \text{P}_4(s) + 6\text{H}_2(g) \rightarrow 4\text{PH}_3(g) \)
   How many liters of phosphine are formed when 0.42 L of hydrogen reacts with phosphorus?

Did you notice that in Sample Problem 12.5 the 22.4 L/mol factors canceled out? This will always be true in a volume-volume problem. Remember that coefficients in a balanced chemical equation indicate the relative numbers of moles. The coefficients also indicate the relative volumes of interacting gases.

Section 12.2 Chemical Calculations 365
### Sample Problem 12.6

**Answers**

19. 18.6 mL SO₂
20. 1.9 dL CO₂

**Math Handbook**

For a math refresher and practice, direct students to dimensional analysis, page R66.

### E ASSESS

#### Evaluate Understanding

**Problem-Solving 12.19** Solve Problem 19 with the help of an interactive guided tutorial.

**Math Handbook**

For help with dimensional analysis, go to page R66.

#### Reteach

Use molecular models to review the importance of mole ratios. Illustrate how the mole ratios from the balanced chemical equation are related to the individual atoms, formula units, and molecules of the reactants and products as described by the equation.

**Connecting Concepts**

The combustion of acetylene gas is represented by this equation:

\[ 2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(g) \]

How many grams of CO₂ and grams of H₂O are produced when 52.0 g C₂H₂ burns in oxygen?

### 12.2 Section Assessment

21. **Key Concept** How are mole ratios used in chemical calculations?
22. **Key Concept** Outline the sequence of steps needed to solve a typical stoichiometric problem.
23. Write the 12 mole ratios that can be derived from the equation for the combustion of isopropyl alcohol.
24. The combustion of acetylene gas is represented by this equation:

**Chemical Quantities** Review the “mole road map” at the end of Section 10.2. Explain how this road map ties into the summary of steps for stoichiometric problems shown in Figure 12.8.

**Connecting Concepts**

**Assessment 12.2** Test yourself on the concepts in Section 12.2.

### Practice Problems

- **19.** Calculate the volume of sulfur dioxide produced when 27.9 mL O₂ reacts with carbon disulfide.
- **20.** How many deciliters of carbon dioxide are produced when 0.38 L SO₂ is formed?

**Section 12.2 Assessment**

21. Mole ratios are written using the coefficients from a balanced chemical equation. They are used to relate moles of reactants and products in stoichiometric calculations.
22. Convert the given quantity to moles; use the mole ratio from the equations to find the moles of the wanted; convert moles of wanted to the desired unit.
23. Evaluate Understanding

**Problem-Solving 12.19** Solve Problem 19 with the help of an interactive guided tutorial.

**Math Handbook**

For help with dimensional analysis, go to page R66.

### Evaluate Does the result make sense?

Because the volume ratio is 2 volumes SO₂ to 1 volume O₂, the volume of O₂ should be half the volume of SO₂. The answer should have three significant figures.

20.4 mL SO₂ \( \times \frac{1 \text{ mL O}_2}{2 \text{ mL SO}_2} = 10.2 \text{ mL O}_2 \)

**Chemical Quantities** Review the “mole road map” at the end of Section 10.2. Explain how this road map ties into the summary of steps for stoichiometric problems shown in Figure 12.8.

**Connecting Concepts**

**Assessment 12.2** Test yourself on the concepts in Section 12.2.

### Practice Problems

- **19.** Calculate the volume of sulfur dioxide produced when 27.9 mL O₂ reacts with carbon disulfide.
- **20.** How many deciliters of carbon dioxide are produced when 0.38 L SO₂ is formed?
Analysis of Baking Soda

**Purpose**
To determine the mass of sodium hydrogen carbonate in a sample of baking soda using stoichiometry.

**Materials**
- baking soda
- 3 plastic cups
- soda straw
- balance
- pipets of HCl, NaOH, and thymol blue
- pH sensor (optional)

**Procedure**

**Small-Scale Lab Manual.**

A. Measure the mass of a clean, dry plastic cup.

B. Using the straw as a scoop, fill one end with baking soda to a depth of about 1 cm. Add the sample to the cup and measure its mass again.

C. Place two HCl pipets that are about 3/4 full into a clean cup and measure the mass of the system.

D. Transfer the contents of both HCl pipets to the cup containing baking soda. Swirl until the fizzing stops. Wait 5–10 minutes to be sure the reaction is complete. Measure the mass of the two empty HCl pipets in their cup again.

E. Add 5 drops of thymol blue to the plastic cup.

F. Place two full NaOH pipets in a clean cup and measure the mass of the system.

G. Add NaOH slowly to the baking soda/HCl mixture until the pink color just disappears. Measure the mass of the NaOH pipets in their cup again.

**Analyze**
Using your experimental data, record the answers to the following questions below your data table.

1. Write a balanced equation for the reaction between baking soda (NaHCO₃) and HCl.
2. Calculate the mass in grams of the baking soda.
3. Calculate the total mmol of 1M HCl.

**You’re the Chemist**
The following small-scale activities allow you to develop your own procedures and analyze the results.

1. Analyze It! For each calculation you did, substitute each quantity (number and unit) into the equation and cancel the units to explain why each step gives the quantity desired.
2. Design It! Baking powder consists of a mixture of baking soda, sodium hydrogen carbonate, and a solid acid, usually calcium dihydrogen phosphate (Ca(H₂PO₄)₂). Design and carry out an experiment to determine the percentage of baking soda in baking powder.

**Expected Outcome**
Sample data: Step A. 2.83 g, B. 3.28 g, C. 10.70 g, D. 4.29 g, E. 10.53 g, G. 8.78 g.