CHAPTER 10: CHEMICAL QUANTITIES

INTRODUCTION

In this chapter you will perform many chemical calculations, all of which are based on fundamental principles, such as balanced equations. A balanced equation can provide more information than is apparent at first glance. You can use a balanced equation to help answer such questions as "How much is produced?" and "How much would be needed to make?" Only a balanced equation will provide correct answers to these questions.

Often, when two reactants are mixed together, one of them will run out before the other one is all used up. In a situation like this, the amount of product you can make will be limited by the reactant that is used up first. The balanced equation will help you determine which reactant runs out, and how much product you can make.

GOALS FOR THIS CHAPTER

1. Understand the different ways to write a balanced chemical equation: with molecules or with moles. (Section 10.1)
2. Know how to use the mole ratio from a balanced chemical equation as a conversion factor to calculate moles of product. (Section 10.2)
3. Know how to use the mole ratio from a balanced chemical equation and the molar mass as conversion factors to calculate the mass of product produced or reactant required. (Section 10.3)
4. Be able to recognize a stoichiometry problem where you need to determine the limiting reactant before you calculate the quantity of product or reactant required. (Section 10.4)
5. Be able to determine which reactant in a chemical reaction is limiting and go from there to calculate the quantity of product produced or reactant required. (Section 10.4)
6. Know how to use the actual yield and the theoretical yield of a chemical reaction to calculate the percent yield. (Section 10.5)

QUICK DEFINITIONS

Mole — 6.022×10^{23} units. A mole of eggs would be 6.022×10^{23} eggs. (Section 10.1)

Mole ratios — A conversion factor which relates the number of moles of product and reactant to each other. In the equation $2H_2 + O_2 \rightarrow 2H_2O$ the mole ratio for hydrogen and oxygen is 2 mol H_2 for each mol of O_2, or $\frac{2 \text{ mol } H_2}{1 \text{ mol } O_2}$. (Section 10.2)

Stoichiometry	The calculation of the relative amounts of reactants or products using a balanced chemical equation. Pronounced stoy ke om etry. (Section 10.3)
Stoichiometric quantities	Amounts of reactants, which, when added together, are both used up, leaving no reactant molecules left over. (Section 10.4)
Limiting reactant	When quantities of two reactants are mixed together in non-stoichiometric quantities, one of the reactants will run out first. This is the limiting reactant. (Section 10.4)
Theoretical yield	The amount of product which could be produced if a chemical reaction were 100% efficient. Theoretical yield is the amount of product you calculate in a stoichiometry problem. (Section 10.5)
Actual yield	The actual amount of product which is produced in a real laboratory situation. Actual yields are smaller than theoretical yields. (Section 10.5)
Percent yield	The mass of product actually obtained compared to the amount which is theoretically possible, expressed as a percent. Large numbers near 100 percent mean a high yield (close to the theoretical) and small numbers mean a low yield (far from the theoretical). (Section 10.5)

PRETEST

1. For the reaction below, express the coefficients in terms of moles.

$$1.506 \times 10^{22} \text{ molecules } CuO(s) \rightarrow 1.506 \times 10^{22} \text{ atoms } Cu(s) + 7.528 \times 10^{21} \text{ molecules } O_2(g)$$

2. How many moles of O_2 are needed to react with 0.035 mol of C_2H_6?

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

3. How many moles of oxygen gas will react with just 0.25 mol N_2 gas?

$$N_2(g) + O_2(g) \rightarrow 2\ NO(g)$$

4. Aluminum metal reacts with hydrochloric acid to produce aluminum chloride and hydrogen gas. How many grams of H_2 gas can be produced from 8.56 g Al?

5. Sulfuric acid reacts with sodium hydroxide to form sodium sulfate and water. How many grams of sodium hydroxide will neutralize 0.838 g H_2SO_4?

6. How much barium sulfate can be formed from the reaction between 0.56 mol barium oxide and 0.35 mol sulfur trioxide?

$$BaO(s) + SO_3(g) \longrightarrow BaSO_4(s)$$

7. How many grams of KNO_3 are produced when 0.982 g $Fe_2(SO_4)_3$ are produced?

$$3\ K_2SO_4(aq) + 2\ Fe(NO_3)_3(aq) \longrightarrow 6\ KNO_3(aq) + Fe_2(SO_4)_3(s)$$

8. When 8.00 g Na_3PO_4 are mixed with 16.0 g $AgNO_3$, which reactant limits the amount of solid Ag_3PO_4 that can be produced?

$$Na_3PO_4(aq) + 3\ AgNO_3(aq) \longrightarrow Ag_3PO_4(s) + 3\ NaNO_3(aq)$$

9. When the quantities of reactants in question 8 are mixed, how many grams of Ag_3PO_4 can be made?

10. When the quantities of reactants in question 8 are mixed, 11.4 g of Ag_3PO_4 are actually made. What is the percent yield?

PRETEST ANSWERS

1. We can convert molecules CuO, atoms of Cu, and molecules of O_2 to moles using Avogadro's number.

$$1.506 \text{ x } \cancel{\text{molecules}} \text{ CuO x } \frac{1 \text{ mol}}{6.022 \text{ x } 10^{23} \ \cancel{\text{molecules}} \text{ CuO}} = 0.02501 \text{ mol Cu O}$$

$$1.506 \text{ x } \cancel{\text{atoms}} \text{ Cu x } \frac{1 \text{ mol}}{6.022 \text{ x } 10^{23} \ \cancel{\text{atoms}} \text{ Cu}} = 0.02501 \text{ mol Cu}$$

$$7.528 \text{ x } 10^{21} \ \cancel{\text{molecules}} \ O_2 \text{ x } \frac{1 \text{ mol}}{6.022 \text{ x } 10^{23} \ \cancel{\text{molecules}} \ O_2} = 0.01250 \text{ mol } O_2$$

We can express our answer as

$$0.02501 \text{ mol CuO} \longrightarrow 0.02501 \text{ mol Cu} + 0.01250 \text{ mol } O_2$$

Its easier to understand the molar relationships between reactants and products if we express the answer in whole numbers.

$$80\,(0.02501 \text{ mol})\ CuO \longrightarrow 80(0.02501 \text{ mol})\ Cu + 80(0.01250 \text{ mol})\ O_2$$

$$2\ CuO \longrightarrow 2\ Cu + O_2 \qquad (10.1)$$

2. We can use the molar ratio to help us find the moles of O_2 needed to react with 0.35 mol C_2H_6

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

Each 2 mol of C_2H_6 requires 7 mol O_2. The molar ratio is $\frac{7 \text{ mol } O_2}{2 \text{ mol } C_2H_6}$.

$$0.35\ \cancel{\text{mol}}\ C_2H_6 \times \frac{7 \text{ mol } O_2}{2\ \cancel{\text{mol}}\ C_2H_6} = 1.2 \text{ mol } O_2 \qquad (10.2)$$

3. The mole ratio for nitrogen and oxygen in this reaction is $\frac{1 \text{ mol } O_2}{1 \text{ mol } N_2}$. Equal numbers of moles of nitrogen and oxygen react, so 0.25 mol nitrogen reacts with just 0.25 mol oxygen.

4. First, write the formulas and physical states for the reactants and products. Then balance the reaction.

$$2\ Al(s) + 6\ HCl(aq) \longrightarrow 2\ AlCl_3(aq) + 3\ H_2(g)$$

We cannot directly convert between grams of Al and grams of H_2. However, any time we are given grams, we can always calculate moles using the molar mass. We will first convert grams of Al to moles of Al. The balanced equation tells us the relationship between moles of Al and moles of H_2, so we can determine the moles of H_2. If we know the moles of H_2, we can use the molar mass to calculate the grams of H_2.

$$8.56\ \cancel{g}\ Al \times \frac{1\ \cancel{\text{mol}}\ Al}{26.98\ \cancel{g}\ Al} \times \frac{3\ \cancel{\text{mol}}\ H_2}{2\ \cancel{\text{mol}}\ Al} \times \frac{2.016 \text{ g } H_2}{1\ \cancel{\text{mol}}\ H_2} = 0.959 \text{ g } H_2$$

grams Al — mol Al — mole ratio — mol H_2 — grams H_2

5. Before attempting to solve this problem, write the balanced equation.

$$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

The problem asks how many grams of sodium hydroxide will neutralize 0.838 g H_2SO_4. This is another way of asking how many grams of sodium hydroxide will react with 0.838 g H_2SO_4. So we can approach this problem as we have several others. First, convert grams of H_2SO_4 to moles of H_2SO_4. Use the mole ratio from the balanced equation to determine the moles of NaOH which will react. Then convert moles of NaOH to grams of NaOH using the molar mass.

$$0.838\ \cancel{g}\ H_2SO_4 \times \frac{1\ m\cancel{o}l\ H_2SO_4}{98.09\ \cancel{g}\ H_2SO_4} \times \frac{2\ m\cancel{o}l\ NaOH}{1\ m\cancel{o}l\ H_2SO_4} \times \frac{40.00\ g\ NaOH}{1\ m\cancel{o}l\ NaOH} = 0.683\ g\ NaOH$$

6. First write the balanced equation for this reaction.

$$BaO(s) + SO_3(g) \longrightarrow BaSO_4(s)$$

We are asked to determine how much product can be made when particular molar quantities of two reactants are mixed. We need to determine which reactant is the limiting reactant. Use the mole ratio from the balanced equation to convert from moles of each reactant to moles of $BaSO_4$.

$$0.56\ m\cancel{o}l\ BaO \times \frac{1\ mol\ BaSO_4}{1\ m\cancel{o}l\ BaO} = 0.56\ mol\ BaSO_4$$

$$0.35\ m\cancel{o}l\ SO_3 \times \frac{1\ mol\ BaSO_4}{1\ m\cancel{o}l\ SO_3} = 0.35\ mol\ BaSO_4$$

The limiting reactant is sulfur trioxide. We know how many moles of $BaSO_4$ can be produced when the given quantities of reactants are mixed, so we can now calculate the grams of barium sulfate from the molar mass of barium sulfate.

$$0.35\ m\cancel{o}l\ BaSO_4 \times \frac{233.4\ g\ BaSO_4}{1\ m\cancel{o}l\ BaSO_4} = 82\ g\ BaSO_4$$

7. We can use the mol ratio that relates moles of KNO_3 and moles of $Fe_2(SO_4)$, $\frac{6\ mol\ KNO_3}{1\ mol\ Fe_2(SO_4)_3}$

$$0.982\ \cancel{g}\ Fe_2(SO_4)_3 \times \frac{1\ m\cancel{o}l\ Fe_2(SO_4)_3}{399.9\ \cancel{g}\ Fe_2(SO_4)_3} \times \frac{6\ m\cancel{o}l\ KNO_3}{1\ m\cancel{o}l\ Fe_2(SO_4)_3} \times \frac{101.11\ g\ KNO_3}{1\ m\cancel{o}l\ KNO_3}$$

$$\times\ 1.49\ g\ KNO_3 \qquad (10.3)$$

8. We need to find the number of moles of Ag_3PO_4 which can be made from 8.00 g Na_3PO_4 and from 16.0 g $AgNO_3$.

$$8.00\ \cancel{g}\ Na_3PO_4 \times \frac{1\ \cancel{mol}\ Na_3PO_4}{163.94\ \cancel{g}\ Na_3PO_4} \times \frac{1\ mol\ Ag_3PO_4}{1\ \cancel{mol}\ Na_3PO_4} = 0.0488\ mol\ Ag_3PO_4$$

$$16.0\ \cancel{g}\ AgNO_3 \times \frac{1\ \cancel{mol}\ AgNO_3}{169.91\ \cancel{g}\ AgNO_3} \times \frac{1\ mol\ Ag_3PO_4}{3\ \cancel{mol}\ AgNO_3} = 0.0314\ mol\ Ag_3PO_4$$

We only have enough $AgNO_3$ to produce 0.031 mol Ag_3PO_4. It is the limiting reactant. (10 .4)

9. $$16.0\ \cancel{g}\ AgNO_3 \times \frac{1\ mol\ A\cancel{g}NO_3}{169.91\ A\cancel{g}NO_3} \times \frac{1\ \cancel{mol}\ Ag_3PO_4}{3\ \cancel{mol}\ AgNO_3} \times \frac{418.67\ g\ Ag_3PO_4}{1\ \cancel{mol}\ Ag_3PO_4} = 13.1\ g\ Ag_3PO_4$$

(10.4)

10. $$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$\frac{11.4\ g}{13.1\ g} \times 100\% = 87.0\%$$ (10.4)

CHAPTER REVIEW

10.1 INFORMATION GIVEN BY CHEMICAL EQUATIONS

How Can You Convert From a Word Description To a Chemical Equation?

An example of a word description of a chemical equation is: When magnesium and oxygen react, magnesium oxide is formed. We can use the information contained in this sentence to help write the chemical equation and to help perform calculations. Because magnesium (Mg) and oxygen (O_2) are mentioned as the substances which are going to react, they are called the reactants and are written on the left side of the chemical equation. The reactants are separated from the products by an arrow

$$Mg + O_2 \longrightarrow$$

The word description says that magnesium oxide (MgO) is formed. In other words, magnesium oxide is the product of the reaction. Products are written on the right side of the equation.

$$Mg + O_2 \longrightarrow MgO$$

We now have a chemical equation that incorporates all of the reactants and products mentioned in the word description. However, a chemical equation must always have the same numbers of each kind of atom on both sides of the equation because atoms are not created or destroyed during a chemical reaction. The reaction between magnesium and oxygen above does not have the same numbers of atoms on both sides. On the left side (reactant side) there are two oxygen atoms and one magnesium atom. On the right side (product side) there is one oxygen atom and one magnesium atom. The equation is **not balanced.** We need to adjust the number of oxygen atoms on the right side of the equation to match the number on the left. If we put a two in front of the formula for magnesium oxide, the number of oxygen atoms on the right matches the number on the left.

$$Mg + O_2 \longrightarrow 2MgO$$

But now, there are two magnesium atoms on the left and one on the right. We can adjust the number on the left to match the number on the right by putting a two in front of the Mg on the left side of the equation.

$$2Mg + O_2 \longrightarrow 2MgO$$

Now, the equation is balanced. When converting a word description of a chemical reaction to a chemical equation, **always** balance the equation after you write the formulas for the reactants and products.

What Information Can You Get From a Balanced Chemical Equation?

Let's look at the reaction between magnesium and oxygen.

$$2Mg + O_2 \longrightarrow 2MgO$$

This chemical equation says that two atoms of magnesium and one molecule of oxygen react to form two molecules of magnesium oxide. If we multiply both sides of a chemical equation by the same number, we do not change the relative numbers of atoms and molecules. If we multiply both sides of the equation by 4,

$$4(2Mg) + 4(O_2) \longrightarrow 4(2MgO)$$

there are now eight atoms of magnesium on the left and eight on the right, and there are eight atoms of oxygen on the left and eight on the right. The equation is still balanced.

$$8Mg + 4O_2 \longrightarrow 8MgO$$

We can multiply both sides of the equation by any number, and the equation will still be balanced. If we multiply both sides of the equation by Avogadro's number (6.022×10^{23}),

$$6.022 \times 10^{23}(2Mg) + 6.022 \times 10^{23}(O_2) \longrightarrow 6.022 \times 10^{23}(2MgO)$$

the equation is still balanced. We still have the same numbers of oxygen and magnesium atoms on both sides of the equation. From your experience with counting by weighing, you know that Avogadro's number of atoms or molecules is equal to one mole of atoms or molecules. So, the chemical equation can be written

$$2 \text{ mole Mg} + 1 \text{ mole } O_2 \longrightarrow 2 \text{ moles MgO}$$

When we write balanced chemical equations like

$$2Mg + O_2 \longrightarrow 2MgO$$

we can interpret them to mean two moles of magnesium react with one mole of oxygen to form two moles of magnesium oxide. When we interpret the coefficients in chemical equations to mean moles instead of individual molecules, we can use the number of moles to calculate the mass of reactants and products involved in the reaction.

10.2 MOLE-MOLE RELATIONSHIPS

How Can You Use a Balanced Equation To Calculate the Number of Moles of Product Produced or Reactant Required?

From the balanced equation for the decomposition of potassium chlorate:

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

two moles of potassium chlorate decompose to produce three moles of oxygen. We can write this as a ratio, called the **molar ratio,** $\frac{2 \text{ mol } KClO_3}{3 \text{ mol } O_2}$. What if a problem asked how many moles of oxygen could be formed from the decomposition of 8.4 moles of potassium chlorate? We can use the balanced equation and a molar ratio to help solve this problem. The given quantity in the problem is moles of potassium chlorate, and the desired quantity is moles of oxygen. There are several molar ratios which can be derived from this balanced equation. The equation says that two moles of $KClO_3$ react to form three moles of O_2 $\frac{2 \text{ mol } KClO_3}{3 \text{ mol } O_2}$, two moles of $KClO_3$ react to form two moles of KCl $\frac{2 \text{ mol } KClO_3}{2 \text{ mol KCl}}$, and two moles of KCl are produced for every three moles of O_2 $\frac{2 \text{ mol KCl}}{3 \text{ mol } O_2}$ The molar ratio we need to solve this problem is $\frac{2 \text{ mol } KClO_3}{3 \text{ mol } O_2}$. By using the molar ratio as a conversion factor, we can cancel the moles of potassium chlorate, and are left with moles of oxygen.

$$8.4\ \cancel{mol}\ KClO_3 \times \frac{3\ mol\ O_2}{2\ \cancel{mol}\ KClO_3} = 13\ mol\ O_2$$

We have used the molar ratio of product to reactant to help determine how many moles of product can be formed from a given number of moles of reactant.

10.3 MASS CALCULATIONS

How Can You Use a Balanced Equation to Calculate the Mass of Reactants or Products?

Chemists are often required to answer practical questions. How much product can be made from a reaction, or how much reactant does it take to make a certain amount of product? To answer such questions, chemists need to know the actual masses of products and reactants, not just the number of moles. With a balanced chemical equation we can answer questions about masses. Section 10.2 showed you how you can answer questions about moles with a balanced equation.

Example:

If you have 125 g of sodium metal, how much chlorine gas is required to react with all of the sodium to produce sodium chloride? First, write the balanced chemical equation for this reaction.

$$2Na + Cl_2 \rightarrow 2NaCl$$

We have seen that the mole ratio is a convenient conversion factor when we are given a number of moles of one component in an equation and we want to calculate the number of moles of some other component. But in this problem we want to know the number of grams, not the number of moles. If we use the atomic mass of sodium, we can find the conversion factor we need and then calculate the number of moles of sodium in 125 g of sodium. One mole of any element is equal to the atomic weight of that element in grams. So one mole of sodium is equal to 22.99 g of sodium. We can use this statement as a conversion factor in the form $\frac{1\ mol\ sodium,}{22.99\ g\ sodium}$ to convert grams to moles.

$$125\ \cancel{g}\ sodium \times \frac{1\ mol\ sodium}{22.99\ \cancel{g}\ sodium} = 5.44\ mol\ sodium$$

Now we know the number of moles of sodium, but we want to know the number of grams of chlorine. We need to select a mole ratio from the balanced equation which relates sodium and chlorine. For every two moles of sodium one mole of chlorine is required, so the conversion factor would be $\frac{1\ mol\ chlorine}{2\ mol\ sodium}$. Sodium appears on the bottom of the conversion factor so that moles of sodium cancel, leaving moles of chlorine.

$$5.44 \text{ } \cancel{\text{mol}} \text{ sodium} \times \frac{1 \text{ mol chlorine}}{2 \text{ } \cancel{\text{mol}} \text{ sodium}} = 2.72 \text{ mol chlorine}$$

We now know the number of moles of chlorine required to react with 125 g of sodium, but the problem asks for the number of grams. One mole of any molecule is equal to its molar mass in grams. A chlorine molecule contains two chlorine atoms so the molar mass of a chlorine molecule is two times the atomic mass of an individual chlorine atom. Two times 35.45 g equals 70.90 g. So one mole of chlorine gas has a molar mass of 70.90 g. Use this statement to construct the conversion factor $\frac{70.90 \text{ g chlorine}}{1 \text{ mol chlorine}}$. Moles of chlorine appear on the bottom of the conversion factor so that they cancel, leaving us with grams of chlorine. To convert moles of chlorine to grams

$$2.72 \text{ } \cancel{\text{mol}} \text{ chlorine} \times \frac{70.90 \text{ g chlorine}}{1 \text{ } \cancel{\text{mol}} \text{ chlorine}} = 193 \text{ g chlorine}$$

Let's see how we can solve this entire problem with one conversion string.

$$125 \text{ } \cancel{\text{g}} \text{ sodium} \times \frac{1 \text{ } \cancel{\text{mol}} \text{ sodium}}{22.99 \text{ } \cancel{\text{g}} \text{ sodium}} \times \frac{1 \text{ } \cancel{\text{mol}} \text{ chlorine}}{2 \text{ } \cancel{\text{mol}} \text{ sodium}} \times \frac{70.90 \text{ g chlorine}}{1 \text{ } \cancel{\text{mol}} \text{ chlorine}} = 193 \text{ g chlorine}$$

Problems such as these, which require the calculation of masses of products or reactants, are often called **stoichiometry** problems.

10.4 CALCULATIONS INVOLVING A LIMITING REACTANT

What Is Meant by the Term Limiting Reactant?

The balanced equation for the reaction between sodium metal and chlorine gas is

$$2Na + Cl_2 \rightarrow 2NaCl$$

When two moles of sodium react with one mole of chlorine, all of the reactants are converted to products. That is, after the reaction is finished, there is no more sodium or chlorine, only sodium chloride. If we add quantities of reactants equal to the mole ratio of the reactants, we will use up all of the reactants. In real chemical experiments, we often do not add enough of each of the reactants to use all of them up. Often, one is used up completely but another one is left over. For example, if we added two moles of sodium to four moles of chlorine, would we use up all the reactants, or would one of them be left over? Let's think through what happens when we mix two moles of sodium with four moles of chlorine. According to the mole ratio, two moles of sodium require one mole of chlorine. Once the two moles of sodium react, there is none left over. When four moles of chlorine react with sodium, only one mole is used up, because there is not enough sodium present to use up all of the chlorine. There are three moles of chlorine left over. The excess chlorine cannot react, because there is no more sodium left. Because the sodium has run out, the reaction stops, even though there is enough chlorine present to make more sodium chloride. In this reaction, sodium is called the **limiting reactant**. Only two moles of sodium chloride will form. In this example, it was straightforward to tell which reactant was the limiting one because all of the quantities were given in moles that you could easily compare. When you are given grams of reactants, you will need to convert to moles to determine which reactant is limiting.

How Can You Determine Which Reactant is Limiting?

Example:

If 36 g sodium reacts with 49 g of chlorine, how much sodium chloride can be produced? In this problem, we cannot tell by looking at the mass of the reactants whether or not they are present in quantities equal to the mole ratio, in other words, whether all of the reactants will be used up, or whether one of them is a limiting reactant. Before we can calculate the grams of sodium chloride produced, we will have to determine whether one of the reactants is limiting. To make the comparison we need to convert grams to moles. Let's see how many moles of product could be formed from each reactant, as we did for the example with moles above. First we will find how many moles of sodium chloride could be produced from 36 g of sodium.

$$36\ \cancel{g}\ \text{sodium} \times \frac{1\ \cancel{\text{mol}}\ \text{sodium}}{22.99\ \cancel{g}\ \text{sodium}} \times \frac{2\ \text{mol sodium chloride}}{2\ \cancel{\text{mol}}\ \text{sodium}} = 1.6\ \text{mol sodium chloride}$$

So, 36 g of sodium, if it were all used, could form 1.6 mol sodium chloride. Now, let's find the number of moles of sodium chloride which could be produced from 49 g of chlorine.

$$49\ \cancel{g}\ \text{chlorine} \times \frac{1\ \cancel{\text{mol}}\ \text{chlorine}}{70.90\ \cancel{g}\ \text{chlorine}} \times \frac{2\ \text{mol sodium chloride}}{1\ \cancel{\text{mol}}\ \text{chlorine}} = 1.4\ \text{mol sodium chloride}$$

If we used up all the sodium, we would form 1.6 mol sodium chloride, and if we used up all the chlorine, we would form 1.4 mol sodium chloride. We cannot form 1.6 mol of sodium chloride because there is only enough chlorine to make 1.4 mol of sodium chloride. Therefore, chlorine is the limiting reactant. It will limit the amount of product we can make. How can we find the mass of sodium chloride? We already know the number of moles we can make. By using the molar mass of sodium chloride as a conversion factor, we can convert moles of sodium chloride to grams of sodium chloride.

$$1.4\ \cancel{\text{mol}}\ \text{sodium chloride} \times \frac{58.44\ \text{g sodium chloride}}{1\ \cancel{\text{mol}}\ \text{sodium chloride}} = 82\ \text{g sodium chloride}$$

When Do You Need to Determine the Limiting Reactant?

When you are given the masses of two reactants and asked how much product you can make, you need to determine which is the limiting reactant before you calculate grams of product.

When you are given the mass of one reactant, but told that the other reactant is present in **excess**, you do not have to calculate the limiting reactant. Because one reactant is present in excess, the other reactant will automatically be the limiting reactant, the one which will run out first.

Example:

When Al reacts with HCl, the products are $AlCl_3$ and H_2.

$$2Al(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3H_2(g)$$

Several different kinds of problems can be written using this balanced equation. For some of them, you will need to calculate the limiting reactant. Let's look at several different problems and determine whether or not it is necessary to find the limiting reactant.

a. When 4.0 g Al react with 5.5 g HCl, how many grams of $AlCl_3$ can be produced? Because masses of two reactants are given, you determine which one is limiting before you can find the mass of product.

b. When 0.3 mol Al reacts with 1.0 mol HCl, how many moles of H_2 can be formed? In this problem you are given moles of two reactants. You must determine which reactant is limiting before you can find moles of product.

c. When 2.5 mol Al reacts with excess HCl, how many moles of $AlCl_3$ can be formed? In this problem you are given the moles of one reactant, but the other reactant is present in excess. In this kind of problem you do not need to determine which is the limiting reactant. Al is limiting. It will run out first because HCl is present in excess.

10.5 PERCENT YIELD

What Is Meant By the Terms Theoretical Yield, Actual Yield and Percent Yield?

Theoretical yield is the amount of product which would be formed under ideal conditions. Theoretical yield is the maximum amount of product possible. Any time we calculate the amount of product which can be formed from a chemical reaction, we are actually calculating the theoretical yield. **Actual yield** is the amount of product you can really make in the laboratory. The actual yield is usually less than the theoretical yield because other reactions which also take place in the reaction container can decrease the amount of product formed. **Percent yield** is a comparison between the actual yield and the theoretical yield. Percent yield is the actual yield divided by the theoretical yield, multiplied by 100 percent.

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\ \%$$

Example:

Magnesium metal reacts with hydrochloric acid to produce hydrogen gas and magnesium chloride.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

When 5.00 g Mg reacts with excess HCl, 0.415 g of H_2 gas can be formed. 0.415 g H_2 is the theoretical yield, the amount we calculate can be formed using the equation below.

$$5.00\ \cancel{g}\ Mg \times \frac{1\ m\cancel{o}l\ Mg}{24.31\ \cancel{g}\ Mg} \times \frac{1\ m\cancel{o}l\ H_2}{1\ m\cancel{o}l\ Mg} \times \frac{2.016\ g\ H_2}{1\ m\cancel{o}l\ H_2} = 0.415\ g\ H_2$$

The actual yield was found in the laboratory to be 0.329 g. What is the percent yield? Percent yield can be calculated by dividing the actual yield by the theoretical yield, and multiplying by 100 %.

$$\text{percent yield} = \frac{0.329\ g\ H_2}{0.415\ g\ H_2} \times 100\ \%$$

$$\text{percent yield} = 79.3\ \%$$

LEARNING REVIEW

1. Rewrite the equation below in terms of moles of reactants and products.

$$6.022 \times 10^{23}\ \text{molecules}\ H_2(g) + 6.022 \times 10^{23}\ \text{molecules}\ I_2(g) \rightarrow 1.204 \times 10^{24}\ \text{molecules}\ HI(g)$$

2. How many moles of hydrogen gas could be produced from 0.8 mol sodium and an excess of water? Solve this problem by writing the equation using moles and by using the mole ratio for sodium and hydrogen.

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

3. How many moles of aluminum oxide could be produced from 0.12 mol Al?

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

4. How many moles of zinc chloride would be formed from the reaction of 1.38 mol Zn with HCl?

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

5. Solid silver carbonate decomposes to produce silver metal, oxygen gas and carbon dioxide.

    a. Write a balanced chemical equation for this reaction.

    b. What mass of silver will be produced by the decomposition of 6.32 g silver carbonate?

6. When aqueous solutions of sodium sulfate and lead(II) nitrate are mixed, a solid white precipitate is formed. How much solid lead(II) sulfate could be produced from 12.0 g Na_2SO_4 if $Pb(NO_3)_2$ is in excess?

$$Na_2SO_4(aq) + Pb(NO_3)_2(aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$$

7. Hydrogen gas and chlorine gas will combine to produce gaseous hydrogen chloride. How many molecules of hydrogen chloride can be produced from 20.1 g hydrogen gas and excess chlorine gas?

8. Some lightweight backpacking stoves use kerosene as a fuel. Kerosene is composed of carbon and hydrogen, and although it is a mixture of molecules, we can represent the formula of kerosene as $C_{11}H_{24}$. When a kerosene stove is lit, the fuel reacts with oxygen in the air to produce carbon dioxide gas and water vapor. If it takes 15 g of kerosene to fry a trout for dinner, how many grams of water are produced?

$$C_{11}H_{24}(l) + 17O_2(g) \rightarrow 11CO_2(g) + 12H_2O(g)$$

9. You are trying to prepare 6 copies of a three-page report. If you have on hand 6 copies of pages one and two, and 4 copies of page three,

    a. How many complete reports can you produce?

    b. Which page limits the number of complete reports you can produce?

10. Manganese(IV) oxide reacts with hydrochloric acid to produce chlorine gas, manganese(II) chloride and water.

$$MnO_2(s) + 4HCl(aq) \rightarrow Cl_2(g) + MnCl_2(aq) + 2H_2O(l)$$

a. When 10.2 g MnO_2 react with 18.3 g HCl, which is the limiting reactant?

b. What mass of chlorine gas can be produced?

c. How many molecules of water can be produced?

11. The acid-base reaction between phosphoric acid and magnesium hydroxide produces solid magnesium phosphate and liquid water. If 121.0 g of phosphoric acid react with 89.70 g magnesium hydroxide, how many grams of magnesium phosphate will be produced?

12. If 85.6 g of potassium iodide react with 2.41 x 10^{24} molecules of chlorine gas, how many grams of iodine can be produced?

$$Cl2(g) + 2KI(s) \longrightarrow 2KCl(s) + I_2(s)$$

13. Aqueous sodium iodide reacts with aqueous lead(II) nitrate to produce the yellow precipitate lead(II) iodide and aqueous sodium nitrate.

a. What is the theoretical yield of lead iodide if 125.5 g of sodium iodide react with 205.6 g of lead nitrate?

b. If the actual yield from this reaction is 197.5 g lead iodide, what is the percent yield?

ANSWERS TO LEARNING REVIEW

1. 6.022 x 10^{23} molecules is equivalent to 1 mol of molecules and 1.204 x 10^{24} molecules is equivalent to 2(6.022 x 10^{23} molecules), so the equation can be rewritten as

$$1 \text{ mol } H_2(g) + 1 \text{ mol } I_2(g) \longrightarrow 2 \text{ mol } HI(g)$$

2. The balanced equation tells us that two moles of sodium react with two moles of water to form two moles of sodium hydroxide and four moles of hydrogen. By using mole ratios determined from the balanced equation, we can calculate the number of moles of reactants required and products produced from 0.8 mol sodium.

$0.8 \cancel{\text{mol}} \text{ Na} \times \dfrac{2 \text{ mol } H_2O}{2 \cancel{\text{mol}} \text{ Na}} = 0.8 \text{ mol } H_2O$ 0.8 mol sodium requires 0.8 mol H_2O.

$0.8 \cancel{\text{mol}} \text{ Na} \times \dfrac{2 \text{ mol NaOH}}{2 \cancel{\text{mol}} \text{ Na}} = 0.8 \text{ mol NaOH}$ 0.8 mol Na produces 0.8 mol NaOH.

$0.8 \cancel{\text{mol}} \text{ Na} \times \dfrac{1 \text{ mol } H_2}{2 \cancel{\text{mol}} \text{ Na}} = 0.4 \text{ mol } H_2$ 0.8 mol Na produces 0.4 mol H_2.