Copy

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

Here the pure metal Mg is oxidized to Mg^{2+}
agent was the hydrochloric acid, but here w
gen in the HCl that is reduced. In HCl, hydrogen's oxidation state ...
hydrogen has been reduced.

The number of different redox reactions is beyond telling. We will look at a few more of them later on.

7.3 Stoichiometry

Everyone who has ever taken a chemistry class remembers performing stoichiometric calculations. The term *stoichiometry* comes from the Greek words *stoicheion*, meaning "element," and *metron*, meaning "measure." When we do stoichiometry, we are measuring the elements. There are many different basic types of calculations we perform in chemistry, but stoichiometric calculations might just be the most basic of the basic. And here's the good news: even though this section has a long, strange foreign name, the calculations themselves are *easy*.

You really have to know only four things to do stoichiometric calculations. First, you need to be able to set up and solve a proportion, a basic math skill you no doubt learned how to do years ago in Prealgebra. Second, you need to be able to compute the molar mass of a compound. This topic was covered in Chapter 2 and should not be difficult at this point. Third, you need to be able to convert from moles to grams and vice versa. This, too, was covered in Chapter 2 and is not difficult. Finally, you need to be able to set up and balance a chemical equation, which we have just covered and which you have (no doubt) just mastered.

So now, with your confidence meter reading "high," let us proceed.

7.3.1 Stoichiometric Calculations

Stoichiometry is all about calculating the quantities of compounds involved in a chemical reaction. For example, just look back at previous page at the reaction of C_3H_8 and O_2. If we have, say, 10.0 kg of propane, how much oxygen will be required to burn it all? How much CO_2 and water will be formed in the process? These calculations are standard stoichiometry.

Let's begin with RULE NUMBER 1.

RULE NUMBER 1
Perform stoichiometric calculations in moles.

If the problem asks for masses in grams, or supplies you with masses in grams, that's fine. If your given quantities are masses in grams, you simply convert given masses into number of moles, as you already know how to do. If you are required to state your answer as a mass in grams, you just convert to grams at the end. But the stoichiometric calculation must performed in moles.

Now for RULE NUMBER 2.

RULE NUMBER 2
The mole ratios for performing the stoichiometric calculations come from the coefficients in the balanced chemical equation.

That's right. That's why this is easy. Remember that while we are balancing a chemical equation, we think of the coefficients as helping us to figure *numbers of atoms*. But remember also that earlier in this chapter, in Section 7.1.4 on balancing chemical equations, I wrote that chemical reactions happen with bulk quantities of atoms, and in chemistry our favorite bulk quantity is moles. So the coefficients in the balanced equation can also be thought of as telling us the *numbers of moles* of compounds or elements that will participate in the reaction.

Now, thinking in moles, look again at that propane equation two pages back. The coefficients in the equation say this: *one* mole of propane will react with *five* moles of oxygen to produce *three* moles of carbon dioxide and *four* moles of water. That's how you do it. The rest is just the details, which I will show you in a few examples.

Before we start in on the examples, let's briefly consider how to use *mole ratios* to perform calculations when the quantities you are given do not match the coefficients. (They never will.) Looking again at the propane equation, the equation says that for every five moles of oxygen involved, four moles of water will form. If you are given a quantity of oxygen and asked to determine how much water will form, you simply use the 5 : 4 ratio as a conversion factor to figure it out. We can write this conversion factor two ways:

$$\frac{5 \text{ mol } O_2}{4 \text{ mol } H_2O} \quad \text{or} \quad \frac{4 \text{ mol } H_2O}{5 \text{ mol } O_2}$$

Now here's how you would use one of these ratios to solve a problem. Suppose the problem is to determine how many moles of water will be produced by the reaction of 13.55 mol O_2 with propane, assuming an unlimited supply of propane. Just take your given quantity and multiply it by the mole ratio that will appropriately cancel with the given quantity to give you the quantity you need:

$$13.55 \text{ mol } O_2 \cdot \frac{4 \text{ mol } H_2O}{5 \text{ mol } O_2} = \frac{13.55 \cdot 4}{5} \text{ mol } H_2O = 10.84 \text{ mol } H_2O$$

Like I said, easy. Now two more comments before we proceed. First, it should be pretty obvious to you that I wrote the mole ratio in that computation as I did because I was given moles of oxygen and I wanted that to cancel to give me moles of water. That's why the oxygen was on the bottom in the conversion factor. If I had been given moles of water and asked to find moles of oxygen, I would have flipped the conversion factor the other way. Second, the coefficients in the chemical equation are *exact*; they are not approximations. When the equation says 5 mol O_2 will react to produce 4 mol H_2O, these figures are exact. Accordingly, they play no part in determining the significant digits to use in your results. In the computation above, the given quantity of oxygen had four significant digits. That's why I wrote the result with four significant digits.

Now we are ready to dive into some examples. The following examples should illustrate everything you need to know. Keep Rule Number 1 in mind: perform the computations in moles. If you are given a quantity as a mass, use the compound's molar mass to convert it to number of moles for the calculation. If you are asked for a result to be stated as a mass, use the molar mass to convert it at the end after you have determined the result in moles.

▼ Example 7.7

In submarines and spacecraft, the breathing air has to be "scrubbed" of excess carbon dioxide to remove the carbon dioxide continuously being exhaled by those on board. One way to purify the breathing air of excess CO_2 is with the compound lithium hydroxide, shown in Figure 7-20. Solid lithium hydroxide reacts with carbon dioxide to produce solid lithium carbonate and water. If the average amount of carbon dioxide exhaled each day is 880 g/person, determine the number of moles of lithium hydroxide that will be consumed per person per day in the reaction.

First, the given quantity is in grams, so we convert this to number of moles. Referring to the periodic table, we determine the molar mass of CO_2, which is 44.010 g/mol. Converting the given quantity to moles, we have

$$880\ \text{g CO}_2 \cdot \frac{1\ \text{mol}}{44.010\ \text{g}} = 20.0\ \text{mol CO}_2$$

Figure 7-20. Lithium hydroxide.

We know our final answer must be stated with two significant digits because of the given quantity. I have kept an extra significant digit in the gram/mole conversion because it is an intermediate result. Final rounding will occur at the end.

Next, write the formula equation:

$$\text{LiOH}\ (s) + \text{CO}_2\ (g) \rightarrow \text{Li}_2\text{CO}_3\ (s) + \text{H}_2\text{O}\ (l)$$

Then, balance the equation.

$$2\text{LiOH}\ (s) + \text{CO}_2\ (g) \rightarrow \text{Li}_2\text{CO}_3\ (s) + \text{H}_2\text{O}\ (l)$$

Now we are ready for the stoichiometry. In the problem, we were given a quantity of CO_2 and asked to find the corresponding quantity of LiOH in the reaction. From the coefficients in the balanced equation, we see that the mole ratio of LiOH to CO_2 in the reaction is 2 : 1. So we use this mole ratio as a conversion factor to convert the mole quantity of CO_2 into a corresponding mole quantity of LiOH.

$$20.0\ \text{mol CO}_2 \cdot \frac{2\ \text{mol LiOH}}{1\ \text{mol CO}_2} = 40.0\ \text{mol LiOH}$$

This is our final result, but it is stated with three significant digits. However, we require two significant digits in our result, and the only way to round 40.0 down to two significant digits is to write it in scientific notation. Thus, our answer is

4.0×10^1 mol LiOH

▲

▼ Example 7.8

One of the body's metabolic processes—from which we get our fuel—is the oxidation of solid glucose by reaction with the oxygen we get from breathing.[2] Glucose, $C_6H_{12}O_6$, is one of the fundamental carbohydrates we take in from eating starch foods such as potatoes and rice. The products of the oxidation reaction are carbon dioxide, which we exhale, and water. (We eliminate more water than we take in; the extra water comes from the oxidation of glucose.) Determine the mass of water produced by the oxidation of 2.00 g glucose.

Again, we are given a quantity in grams, so we will convert this to moles right at the beginning. Using data from the periodic table, the molar mass of glucose is found to be 180.157 g/mol. Converting the given mass of glucose into moles, we have

2 Here is another instance of using the term *oxidation* in reference to an entire molecule. If you look at the oxidation states, you will see that it is the carbon in the glucose that is being oxidized.

$$2.00 \text{ g } C_6H_{12}O_6 \cdot \frac{1 \text{ mol } C_6H_{12}O_6}{180.157 \text{ g } C_6H_{12}O_6} = 0.01110 \text{ mol } C_6H_{12}O_6$$

Again, I have kept an extra significant digit in this intermediate calculation to avoid the build up of rounding error. I will round to the required three significant digits at the end. Next, we write the formula equation for the oxidation of glucose:

$$C_6H_{12}O_6\ (s) + O_2\ (g) \rightarrow CO_2\ (g) + H_2O\ (l)$$

Now balance the equation:

$$C_6H_{12}O_6\ (s) + 6O_2\ (g) \rightarrow 6CO_2\ (g) + 6H_2O\ (l)$$

This problem gives us a quantity of glucose and asks about a quantity of water, so these are the two compounds to look at in the equation. From the coefficients in the balanced equation, we see that the mole ratio of glucose to water in this reaction is 1 : 6. Now write down the given quantity of glucose (in moles) and use the mole ratio to convert this into a mole quantity of water.

$$0.01110 \text{ mol } C_6H_{12}O_6 \cdot \frac{6 \text{ mol } H_2O}{1 \text{ mol } C_6H_{12}O_6} = 0.06660 \text{ mol } H_2O$$

Now that we have the amount of water produced, we simply need to use the molar mass of water to convert this into grams as the problem requires. From the periodic table, we determine that the molar mass of water is 18.02 g/mol. Using this value as a conversion factor, we determine the mass of water produced by the reaction:

$$0.06660 \text{ mol } H_2O \cdot \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 0.06660 \cdot 18.02 \text{ g } H_2O = 1.20 \text{ g } H_2O$$

Notice, finally, that this result has been rounded to three significant digits as the problem requires.

▲

It should be apparent from these two examples that a stoichiometric calculation can be performed using any or all of the products or reactants in a reaction. It doesn't matter which ones are asked about in the problem; you simply set up mole ratios using two compounds (or elements) at a time. We will use this principle in our final example.

▼ Example 7.9

The photosynthesis reaction in plants is identical to the oxidation of glucose running in reverse. In the plant, CO_2 from the air and water from the soil are converted (using sunlight as an energy source, of course) into glucose, which becomes fuel and building material for the plant, and oxygen, which is released into the atmosphere. For each 5.00 g H_2O consumed by a plant, determine the masses of the CO_2 consumed, the O_2 released, and the glucose produced.

Since this reaction is identical to the one in the previous problem, only running in reverse, we can write down the balanced equation immediately:

$$6CO_2\ (g)\ +\ 6H_2O\ (l)\ \rightarrow\ C_6H_{12}O_6\ (s) + 6O_2\ (g)$$

In this example, we are required to calculate everything, so we may as well calculate all four of the molar masses right now. Since the mass given in the problem has three significant digits, I will write down each of the molar masses with four significant digits.

CO_2:	44.01 g/mol
H_2O:	18.02 g/mol
$C_6H_{12}O_6$:	180.2 g/mol
O_2:	32.00 g/mol

Next, let's convert the given quantity from grams to moles.

$$5.00\ \text{g}\ H_2O \cdot \frac{1\ \text{mol}\ H_2O}{18.02\ \text{g}\ H_2O} = 0.2775\ \text{mol}\ H_2O$$

(This value has the extra significant digit required for an intermediate calculation.) We are given an amount of H_2O in the problem, and from this quantity we must calculate quantities of three other substances. So we need mole ratios for water and the three other substances. From the coefficients in the balanced equation, these ratios are

$H_2O : CO_2$	6 : 6, which is equal to 1 : 1
$H_2O : C_6H_{12}O_6$	6 : 1
$H_2O : O_2$	6 : 6, which is equal to 1 : 1

Now we just use the given amount of water, in moles, with one of these ratios to compute the quantities of the other three substances.

$$0.2775\ \text{mol}\ H_2O \cdot \frac{1\ \text{mol}\ CO_2}{1\ \text{mol}\ H_2O} = 0.2775\ \text{mol}\ CO_2$$

$$0.2775\ \text{mol}\ H_2O \cdot \frac{1\ \text{mol}\ C_6H_{12}O_6}{6\ \text{mol}\ H_2O} = 0.04625\ \text{mol}\ C_6H_{12}O_6$$

$$0.2775\ \text{mol}\ H_2O \cdot \frac{1\ \text{mol}\ O_2}{1\ \text{mol}\ H_2O} = 0.2775\ \text{mol}\ O_2$$

The last step is to use the molar masses to convert each of these mole quantities into grams.

$$0.2775 \text{ mol } CO_2 \cdot \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 12.2 \text{ g } CO_2$$

$$0.04625 \text{ mol } C_6H_{12}O_6 \cdot \frac{180.2 \text{ g } C_6H_{12}O_6}{1 \text{ mol } C_6H_{12}O_6} = 8.33 \text{ g } C_6H_{12}O_6$$

$$0.2775 \text{ mol } O_2 \cdot \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 8.88 \text{ g } O_2$$

These quantities have all been rounded to three significant digits as required. We have completed the problem, and there were a lot of calculations involved. A great way to perform an overall error check on all the calculations is to verify that the law of conservation of mass in chemical reactions has been satisfied. From the equation, the masses of the water and carbon dioxide should add up to equal the masses of the glucose and oxygen.

mass CO_2 + mass H_2O = 12.2 g + 5.00 g = 17.2 g

mass $C_6H_{12}O_6$ + mass O_2 = 8.33 g + 8.88 g = 17.21 g

Rounding that second value to three significant digits, we see that the reaction consumed 17.2 g of reactants and produced 17.2 g of products. Mass conservation has been verified, so it is highly likely that our calculations are all correct.

▲

7.3.2 Limiting Reactant

Let's say you have a portable propane burner and a small canister of propane to fuel it. We know that the combustion of propane is a chemical reaction with oxygen that produces carbon dioxide and water. In the combustion reaction of this canister of propane, what is the limiting factor that determines how much water and carbon dioxide will be produced? By *limiting factor* what I really mean with this question is *limiting reactant*—the reactant that will run out first once the reaction starts.

Clearly, the oxygen available for this reaction, which is in the atmosphere, is available without limit. Thus, the limiting reactant is the propane, and when the propane is completely consumed, the reaction will cease. Every chemical reaction has a limiting reactant. One of the reactants will be in shorter supply than the others and will run out first. When it does, the reaction will cease. The quantities of products produced by the reaction are determined by the quantity of the limiting reactant available for the reaction.

The limiting reactant is also often called the *limiting reagent* (pronounced ree-A-gent), a reagent simply being one of the compounds consumed in the reaction.

If you have known quantities available for a chemical reaction, you can determine which reactant is the limiting reactant simply by looking at the mole quantities available and the mole ratios from the balanced chemical equation. Once you know which compound is the limiting reactant, you can calculate the quantities of the products that will be produced by the reaction based on the quantity of the limiting reactant supplied. As with the stoichiometric calculations we just reviewed, all the calculations associated with determining the limiting reactant must be performed in moles. If you are given mass quantities, begin your solution by converting masses into numbers of moles.

▼ Example 7.10

One of the most important industrial chemical processes in the world is the Haber-Bosch process for producing ammonia from nitrogen in the air. Grains need nitrogen as a nutrient, but cannot get it from the air because the nitrogen molecules are held together by the strong nitrogen triple bond. Bacteria in the soil can produce the "fixed" nitrogen that plants can use, but bacteria cannot supply the nitrogen needed for industrial scale agriculture. The nitrogen-based fertilizers used are made from ammonia. Some have estimated that one third of the world's population is sustained by food grown using fertilizer produced by ammonia made with the Haber-Bosch process.

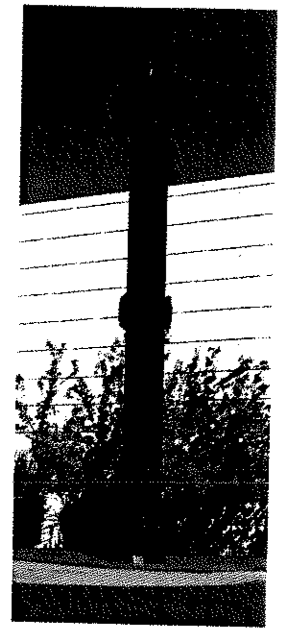

Figure 7-21. This high-pressure reactor from 1921 is on display at the Karlsruhe Institute of Technology in Germany.

The process occurs at extremely high pressures—around 200 atmospheres, or 3,000 psi—in reactors like the historical 1921 reactor shown in Figure 7-21. For years chemists thought the reaction was impossible, but German chemist Fritz Haber solved the problem in 1909, and with engineering assistance from Carl Bosch, had the commercial production of ammonia up and running by the end of 1910. The reaction is as follows:

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

Consider a test run of this process, with 652 kg N_2 and 175 kg H_2 available for the reaction. Determine whether the N_2 or the H_2 is the limiting reactant. Then assuming that the limiting reactant will be completely consumed in the reaction, calculate the mass of ammonia that will have been produced when the reaction ceases.

We begin by converting our mass quantities from kilograms to grams, and from grams to moles. We use the periodic table to determine the molar masses of H_2 and N_2, and then we use these as conversion factors to get mole quantities:

$$175{,}000\ \text{g}\ H_2 \cdot \frac{1\ \text{mol}\ H_2}{2.016\ \text{g}\ H_2} = 86{,}810\ \text{mol}\ H_2$$

$$652{,}000\ \text{g}\ N_2 \cdot \frac{1\ \text{mol}\ N_2}{28.01\ \text{g}\ N_2} = 23{,}280\ \text{mol}\ N_2$$

These mole quantities have each been written with four significant digits, one more than we will have in the final result. Looking now at the balanced equation for the reaction, we see that the ratio of hydrogen to nitrogen required is 3 : 1. In order to consume the 23,280 moles of nitrogen we have available, we need 3(23,280) = 69,840 moles of hydrogen. We have a lot more hydrogen than this available, so nitrogen is the limiting reagent.

Now we perform a standard stoichiometric calculation to determine the amount of ammonia produced from 23,280 mol N_2. The mole ratio of N_2 to NH_3 is 1 : 2, so we use this ratio as a conversion factor to determine the amount of NH_3 produced:

$$23{,}280\ \text{mol}\ N_2 \cdot \frac{2\ \text{mol}\ NH_3}{1\ \text{mol}\ N_2} = 46{,}560\ \text{mol}\ NH_3$$

Finally, we convert this mole quantity to mass as the problem requires. This requires calculating the molar mass of NH_3 to use as a conversion factor.

$$46{,}560 \text{ mol } NH_3 \cdot \frac{17.03 \text{ g } NH_3}{1 \text{ mol } NH_3} = 792{,}900 \text{ g } NH_3 \cdot \frac{1 \text{ kg}}{1000 \text{ g}} = 793 \text{ kg } NH_3$$

This final result has been rounded to three significant digits as the given information requires.

▲

▼ Example 7.11

A 5.00-g strip of magnesium metal is placed in an aqueous solution containing 15.0 g $ZnCl_2$, causing the following reaction:

$$Mg\ (s) + ZnCl_2\ (aq) \rightarrow Zn\ (s) + MgCl_2\ (aq)$$

Determine (a) the limiting reactant, (b) the mass of $MgCl_2$ produced, and (c) the mass of the excess reactant that will remain.

Converting the given quantities to moles, we have:

$$5.00 \text{ g Mg} \cdot \frac{1 \text{ mol Mg}}{24.305 \text{ g Mg}} = 0.2057 \text{ mol Mg}$$

$$15.0 \text{ g } ZnCl_2 \cdot \frac{1 \text{ mol } ZnCl_2}{136.3 \text{ g } ZnCl_2} = 0.1101 \text{ mol } ZnCl_2$$

The mole ratio for Mg and $ZnCl_2$ in the balanced equation is 1 : 1, so the $ZnCl_2$ will be the limiting reactant.

The mole ratio of $ZnCl_2$ and $MgCl_2$ is also 1 : 1, so 0.1101 mol $ZnCl_2$ will produce 0.1101 mol $MgCl_2$, giving a mass of

$$0.1101 \text{ mol } MgCl_2 \cdot \frac{95.12 \text{ g } MgCl_2}{1 \text{ mol } MgCl_2} = 10.5 \text{ g } MgCl_2$$

Since the mole ratio for Mg and $ZnCl_2$ is 1 : 1, 0.1101 mol $ZnCl_2$ will consume 0.1101 mol Mg. The reaction began with 0.2057 mol Mg, so 0.2057 – 0.1101 mol Mg = 0.0956 mol Mg will be left over. This converts to a mass of

$$0.0956 \text{ mol Mg} \cdot \frac{24.305 \text{ g Mg}}{1 \text{ mol Mg}} = 2.32 \text{ g Mg}$$

▲

7.3.3 Theoretical Yield and Percent Yield

In the example of the Haber-Bosch process to produce ammonia, the 793 kg of ammonia we calculated that would be produced by the reaction is called the *theoretical yield*. Let's now assume we perform the reaction using the quantities of reactants listed in that example and we find that we actually end up with 742 kg NH_3 when all the N_2 is consumed and the reaction ceases. This quantity is called the *actual yield*. We can use this quantity to calculate the *percent yield* as follows:

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

Accordingly, the percent yield associated with an actual yield of 742 kg NH_3 is:

$$\text{percent yield} = \frac{742 \text{ kg } NH_3}{793 \text{ kg } NH_3} \times 100\% = 93.6\%$$

When chemists develop procedures to synthesize compounds, the percent yield is a measure of the effectiveness of the procedure. Yields above 80% are considered very good.

Chapter 7 Exercises

SECTION 7.1

1. Butane, C_4H_{10}, is the fuel used in disposable lighters. Butane burns in oxygen to produce carbon dioxide and water. Write the formula equation for this reaction and balance it.
2. Iron (II) sulfide reacts with hydrochloric acid to produce hydrogen sulfide and iron (II) chloride. Write the formula equation for this reaction and balance it.
3. Acetylene, C_2H_2, is a gas that produces an extremely hot flame when combusting with oxygen, so hot that cutting torches use this gas to cut through steel. The products of the combustion are carbon dioxide and water. Write the formula equation for this reaction and balance it.
4. With the aid of a catalyst, liquid methanol, CH_3OH, can be produced from a reaction of carbon monoxide gas and hydrogen gas. Write the formula equation for this reaction and balance it.
5. Potassium carbonate can be used to neutralize the sulfuric acid in well water caused by acid rain. The products of the neutralization reaction are potassium sulfate, carbon dioxide, and water. Write the formula equation for this reaction and balance it.
6. The tungsten metal used as filaments in light bulbs is produced by reacting tungsten oxide with hydrogen gas to produce tungsten metal and water. Write the formula equation for this reaction and balance it.
7. Sodium azide, NaN_3, is the original compound used in automobile air bags. Other compounds were used to initiate the reaction, but the bags were filled by the conversion of solid sodium azide into sodium metal and nitrogen gas. Write the formula equation for this reaction and balance it.
8. Identify the oxidation state for each element in the following compounds, elements and ions:

a. PBr_3	b. As_2O_5	c. ClO_4^-	d. F_2
e. UF_6	f. CO_3^{2-}	g. H_3PO_4	h. $Zn(NO_3)_2$
i. C_4H_{10}	j. $Cr_2O_7^{2-}$	k. KH	l. Fe

9. Why must chemical equations be balanced?