

## Chapter 9

## STOICHIOMETRY, Part 3

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At this point we get into the use of volume in stoichiometry problems. All of our discussions and examples up to now have focused on mass, but volume quantities are also very common. I mentioned this in the beginning of Chapter 7.

“ Since mass and volume are measurable things, they can provide the necessary connection to the number of formula units in a real sample. Then, when it comes time to do a reaction with some number of formula units, we can calculate the mass or the volume which is needed to provide the desired number for each reactant. We can also calculate the mass or volume for each product in the process. These mass and/or volume relationships between the reactants and products of a reaction are called "stoichiometry". ”

Everything still boils down to the critical need to be able to measure how many moles are in a given amount of a reagent. Mass is only one such method. This is the most convenient method when the reagent is a solid. On the other hand, mass is not so convenient when the reagent is a liquid or the reagent is dissolved in a liquid solution. These cases are more easily handled by volumes. For these, we want a connection between volume and moles. These connections typically differ for reagents which are liquids as compared to reagents which are dissolved in liquid solution. I'll do the case of a liquid reagent first since it's easy. Then we'll do reagents dissolved in solution.

## 9.1 Liquid reagents

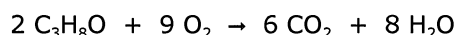
If a reactant or a product is a liquid, then volume is easily accommodated using density. If you're given a volume and a density, this gets you mass automatically. Mass then connects to moles by the usual molar mass. A stoichiometry problem which incorporates density still follows the four basic stoichiometry steps. The only difference is that there will be an extra conversion.

Let's illustrate this.

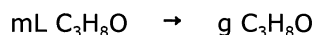
.....  
**Example 1.** How many grams of water can be produced in the combustion of 50.0 mL of isopropyl alcohol? We need the density of C<sub>3</sub>H<sub>8</sub>O: it's 0.786 g/mL.  
 .....

We start right into the standard stoichiometry steps.

• Step 1: Balance the equation. I picked a combustion reaction which we did in Chapter 6. Go back and look at it if you want. The final result was



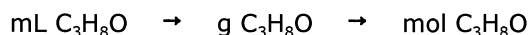
• Step 2 says convert the given amount (50.0 mL) of reagent (isopropyl alcohol) to moles. Well, we don't have a direct volume-to-mole conversion factor, but we do have a volume and a density and that will get us mass. So we first convert volume to grams using density.



Like any conversion factor, density can be used in either of two forms, right-side-up or upside-down.

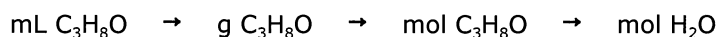
You will need  $\frac{0.786 \text{ g C}_3\text{H}_8\text{O}}{\text{mL C}_3\text{H}_8\text{O}}$  OR  $\frac{\text{mL C}_3\text{H}_8\text{O}}{0.786 \text{ g C}_3\text{H}_8\text{O}}$

You will use the one on the left in the current problem. After that, it's routine stoichiometry. Next, you need molar mass to get to moles.

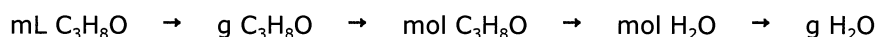


Of the two possible molar mass conversion factors, you need:  $\frac{\text{mol C}_3\text{H}_8\text{O}}{60.09 \text{ g C}_3\text{H}_8\text{O}}$

• Step 3. Rxn ratio, which happens to be 2:8 or 8:2 for this problem. (Which one do you use here?)



• Step 4. Molar mass for H<sub>2</sub>O, 18.02 g. (Which way is it used here: right-side-up or upside-down?)



- Now the big string.

$$\begin{array}{ccccccc}
 \text{mL C}_3\text{H}_8\text{O} & \rightarrow & \text{g C}_3\text{H}_8\text{O} & \rightarrow & \text{mol C}_3\text{H}_8\text{O} & \rightarrow & \text{mol H}_2\text{O} & \rightarrow & \text{g H}_2\text{O} \\
 50.0 \text{ mL C}_3\text{H}_8\text{O} & \times & \frac{0.786 \text{ g C}_3\text{H}_8\text{O}}{\text{mL C}_3\text{H}_8\text{O}} & \times & \frac{\text{mol C}_3\text{H}_8\text{O}}{60.09 \text{ g C}_3\text{H}_8\text{O}} & \times & \frac{8 \text{ mol H}_2\text{O}}{2 \text{ mol C}_3\text{H}_8\text{O}} & \times & \frac{18.02 \text{ g H}_2\text{O}}{\text{mol H}_2\text{O}}
 \end{array}$$

Plug it in, punch it out, round it off: you get 47.1 g H<sub>2</sub>O. That's how much water can be made from the combustion of 50.0 mL of isopropyl alcohol.

This Example nicely illustrates that Step 2 may sometimes need an extra conversion. Actually, this can also happen in Step 4 if your problem asks for the answer in volume; this would also take a density conversion factor. I mentioned the possibility of extra conversions back in Chapter 7 when I first introduced the stoichiometry steps. This was our first example. Be aware that these things can happen.

That concludes the example for liquid reagents. We now turn to stoichiometry involving reagents dissolved in a solution.

## 9.2 The solution to solutions

Zillions of reactions occur in solution on a daily basis. Just ask your own cells. Or check out the ocean or a lake or even a drop of rain. Nature isn't the only one to use solution chemistry, since industry also conducts a multitude of reactions in solution. Here, I'm referring to liquid solutions but there are also gas and solid solutions. The most common gas solution is Earth's atmosphere. It's composed of a number of gases, mostly O<sub>2</sub> and N<sub>2</sub>, but also small or trace amounts of Ar, CO<sub>2</sub>, Ne, CH<sub>4</sub> and other things good and bad. Gas solutions are another very important medium for doing chemical reactions, and our H<sub>2</sub>/Cl<sub>2</sub> reaction in Chapter 7 was of this type. On the other hand, solid solutions do exist, but they are not a common medium for doing reactions at typical conditions.

For now, we focus on liquid phase solutions. Such solutions are composed of two types of components. The first type is the solvent and the second type includes one or more solutes. The solutes are the things that actually dissolve in the solvent; the solute is broken up into individual molecules or ions, mixed in with solvent molecules. Floaters are not solutes. Cloudy things are not solutes. Floaters and cloudy things actually involve particles of one substance mixed in with another substance, although the particle sizes may be too small to see with just the eyeball. A true solution is completely clear. Colors are OK for solutions but cloudiness is not. Clear blue. Clear red. Clear chartreuse. Clear colorless. A clear solution occurs when everything is dissolved, regardless of color. Clouds and floaters are not dissolved and these mixtures are not solutions. Coffee by itself is a solution (assuming you filtered it properly). Coffee plus creamer is not a solution. But I don't use creamer anyway.

By the way, a useful adjective for a true solution is "homogeneous". A homogeneous solution is clear. The otherwise is "heterogeneous", which applies to a mixture in which one or more components is not truly dissolved. These words also apply for gas and solid mixtures.

Many different kinds of substances can be solutes. It doesn't matter if the solute by itself is a solid, liquid or gas. It only matters whether it can dissolve in the solvent. Sodium chloride, NaCl, by itself is a solid. It dissolves in water to form a salt solution. Ethylene glycol, C<sub>2</sub>H<sub>6</sub>O<sub>2</sub>, by itself is a liquid. It's the most common anti-freeze and anti-boil component used in automotive radiators. The compound dissolves easily in water and these solutions are keeping your engine at a reasonable temperature. (Ethylene glycol is colorless. Radiator antifreezes come in colors, like green, orange and even fluorescents, but the colors are due to other solutes which are present.) Gases can dissolve in water and one of the most important examples on the planet is O<sub>2</sub>. Just ask a fish: they breathe the dissolved O<sub>2</sub> right out of the water.

Notice that I have been emphasizing aqueous solutions. This is a well deserved emphasis. The most important solvent on this planet is water. No other solvent comes close. I don't just mean life, either. Reactions in water help shape the planet over the eons.

Many other solvents do exist, however, and many find use when reactions are done in the lab or in industry. Alcohols are common solvents. The term "alcohol" refers to a family of compounds, some of which you may be aware of. The simplest is methyl alcohol (also called methanol), CH<sub>3</sub>OH. You can buy this stuff in stores where it's sold as a common fuel-line antifreeze and as a solvent in some paint departments. Read the label; it's on there. Ethyl alcohol (also called ethanol) is the common grain or drinking alcohol, C<sub>2</sub>H<sub>5</sub>OH. It's also used as a solvent, but for these applications they often "denature" it, which means they add a poison to it. Why? Ethyl alcohol for drinking purposes is regulated and taxed as such; denatured ethyl alcohol for use as a solvent is not regulated and not taxed to the same extent, and they don't want you drinking the less expensive stuff. One of the common poisons they add is methyl

alcohol and the consumption of methyl alcohol can easily cause blindness. Another example of an alcohol is the one we worked with earlier, isopropyl alcohol ( $\text{CH}_3\text{CHOHCH}_3$  or just  $\text{C}_3\text{H}_8\text{O}$ ). This is also available as a fuel-line antifreeze, but it's more commonly known for its presence in rubbing alcohol. In the case of rubbing alcohol, water is the solute and the alcohol is the solvent.

Alcohols are examples of "nonaqueous" solvents. A nonaqueous solvent is any solvent other than water. Other examples are probably in your home. Paint thinner is a solvent which is a mixture of petroleum compounds. Nail polish remover contains acetone or ethyl acetate as the solvent which actually dissolves the polish. Read the label; it's on there. These things are part of your world.

Let's proceed to quantitative aspects involving solutions. First, let me point out that you can still do grams when working with a reaction in solution. Our example last Chapter was exactly of this type when we talked about the reaction between silver nitrate and sodium bromide. In that example, we started with a measured mass of each reactant by itself. This differs from where we are going now. Now we will deal with a measured volume of a solution of the reagent.

The key to this approach is concentration. Concentration can be expressed in many different ways. Rubbing alcohol is commonly 70%  $\text{C}_3\text{H}_8\text{O}$  along with some  $\text{H}_2\text{O}$ . That's a concentration. The hydrogen peroxide they sell in stores is 3%  $\text{H}_2\text{O}_2$  dissolved in  $\text{H}_2\text{O}$ . That's a concentration. Both of these are examples of percent concentration. Percent concentrations are based on mass and/or volume relationships between a solute and the grand sum of all components in the solution.

$$\text{percent concentration} = \frac{\text{mass-or-volume of solute}}{\text{sum of masses-or-volumes of all components}} \times 100\%$$

You can work with these in stoichiometry; I won't go into that right now, but we will see an example in Chapter 15 and we will do a bit more with percents in Chapter 42. Put a star in the margin so I can refer you back here later. Your instructor may include these calculations at this time.

What I want to do right now is connect volume of the solution to moles of the solute. And we need a direct connection, not some beat-around-the-bush percentage. What we want is a concentration term which looks like the following.

$$\frac{\text{moles of solute}}{\text{volume of solution}}$$

And that's exactly what we have: it's called "molarity" and it's abbreviated  $M$ . It's the moles of solute per liter of solution.

$$\text{molarity} = M = \frac{\text{mol solute}}{\text{L soln}}$$

I just snuck in another abbreviation on you: soln. "Soln" is a very common abbreviation for solution. I'll be using it a lot.

There's some terminology and symbolism here to be careful with. Molarity is one kind of concentration expression. The unit is "molar". (Yes, spelled and pronounced just like the tooth.) The "molar" unit is abbreviated  $M$  and it is equal to one mole of solute per liter of solution. Brackets, [like these], are used as a symbol for molarity. The solute formula is placed inside the brackets. Let me show you how some of this goes. Let's say you have 1.0 mol of  $\text{NaCl}$  and you dissolve it in water. You use just enough water so that the final solution volume is 1.0 L after everything is dissolved. We can say this or write this in any of the following ways. All of these statements say the same thing.

The concentration of  $\text{NaCl}$  is 1.0 molar or 1.0  $M$ .

The molarity of  $\text{NaCl}$  is 1.0 molar or 1.0  $M$ .

The solution is 1.0  $M$   $\text{NaCl}$ .

$M$  of  $\text{NaCl} = 1.0 M$

$[\text{NaCl}] = 1.0 M$

Notice that I said you use just enough water to make the final volume 1.0 L. You actually don't have to measure the volume of the solvent by itself. It doesn't matter because the solute can affect the volume anyway. It's very common for the solution volume to be different from the solvent volume. The only thing that matters for molarity is the final volume of the whole solution. Another point to note is that we are talking concentration, not the actual number of moles by itself and not the actual number of liters by itself. Molarity is the ratio of the two. If I have 0.50 mol  $\text{NaCl}$  dissolved in 0.50 L of solution, then all the above statements still apply.

We will do two examples with a basic molarity calculation before proceeding to stoichiometry.

**Example 2.** You have an aqueous solution of methyl alcohol which contains 14.2 g CH<sub>3</sub>OH in 0.2500 L solution. What is the molarity of CH<sub>3</sub>OH in this solution?

The problem asks for a concentration in terms of moles CH<sub>3</sub>OH per liter of solution. Notice that you were given a concentration term already, but it's in the wrong units: you are given a concentration in terms of grams CH<sub>3</sub>OH per liter of solution. We need to convert this to moles per liter.

$$\text{Given: } \frac{14.2 \text{ g CH}_3\text{OH}}{0.2500 \text{ L soln}} \qquad \text{Need: } \frac{??? \text{ mol CH}_3\text{OH}}{\text{L soln}}$$

The denominator has the correct unit, L soln. It's the numerator that needs to change. How? Molar mass. All you need is a molar mass conversion factor for the calculation. Start with the given info.

$$\frac{\text{g CH}_3\text{OH}}{\text{L soln}}$$

The molar mass for CH<sub>3</sub>OH is 32.04 g. Use this to convert the numerator from g to mol.

$$\frac{\text{g CH}_3\text{OH}}{\text{L soln}} \rightarrow \frac{\text{mol CH}_3\text{OH}}{\text{L soln}}$$

Since the denominator started with the desired volume unit (L soln), leave it alone. Bring in all the values.

$$\frac{\text{g CH}_3\text{OH}}{\text{L soln}} \rightarrow \frac{\text{mol CH}_3\text{OH}}{\text{L soln}}$$

$$\frac{14.2 \text{ g CH}_3\text{OH}}{0.2500 \text{ L soln}} \times \frac{\text{mol CH}_3\text{OH}}{32.04 \text{ g CH}_3\text{OH}} = 1.77 \frac{\text{mol CH}_3\text{OH}}{\text{L soln}}$$

This says that the molarity of CH<sub>3</sub>OH in the solution is 1.77 M. We can also write [CH<sub>3</sub>OH] = 1.77 M.

A lot of molarity problems use mL instead of L directly. Let's say the problem above was instead given as 14.2 g CH<sub>3</sub>OH dissolved in 250.0 mL solution. How would that change things? There are two ways of dealing with this. One is to incorporate the power-of-ten prefix conversion into the string. Let's restart the problem, this time with mL in the denominator.

$$\frac{\text{g CH}_3\text{OH}}{\text{mL soln}}$$

As above, use molar mass to convert the numerator from g to mol,

$$\frac{\text{g CH}_3\text{OH}}{\text{mL soln}} \rightarrow \frac{\text{mol CH}_3\text{OH}}{\text{mL soln}}$$

but now we have to change the denominator in another step.

$$\frac{\text{g CH}_3\text{OH}}{\text{mL soln}} \rightarrow \frac{\text{mol CH}_3\text{OH}}{\text{mL soln}} \rightarrow \frac{\text{mol CH}_3\text{OH}}{\text{L soln}}$$

All of this means that our string has one more step than above. No big deal. Bring in the numbers.

$$\frac{\text{g CH}_3\text{OH}}{\text{mL soln}} \rightarrow \frac{\text{mol CH}_3\text{OH}}{\text{mL soln}} \rightarrow \frac{\text{mol CH}_3\text{OH}}{\text{L soln}}$$

$$\frac{14.2 \text{ g CH}_3\text{OH}}{250.0 \text{ mL soln}} \times \frac{\text{mol CH}_3\text{OH}}{32.04 \text{ g CH}_3\text{OH}} \times \frac{10^3 \text{ mL soln}}{\text{L soln}} = 1.77 \frac{\text{mol CH}_3\text{OH}}{\text{L soln}}$$

This brings us again to [CH<sub>3</sub>OH] = 1.77 M.

The second way to deal with mL is the easier way: just do a direct decimal shift on the number itself. Milli- is a three-place decimal shift, so 250.0 mL = 0.2500 L. Just do this and you won't have to mess with a separate conversion step. Some of the other powers-of-ten prefixes might not work so well. In my research group we use microliters of reagents a lot. Decimal shifting six places is prone to a mistake if you're not careful.

Next Example.

**Example 3.** You need to prepare 100.0 mL of an aqueous solution of 2.19 M sodium chloride. How many grams of NaCl will you need to make this solution?

To do this, recall the definition for molarity.

$$\text{molarity} = M = \frac{\text{mol solute}}{\text{L soln}}$$

There are three terms in this relationship: molarity, moles of solute and liters of solution. Anytime you have two of the terms, the third can be calculated. REMEMBER THIS. The present problem provides molarity and it provides volume of solution; that means you can find moles of solute.

$$\text{L soln} \times \text{molarity} = \text{mol solute}$$

That gives moles of solute, but the problem asks for grams of solute. Then what? How do you get from moles to grams? You don't even have to ask: molar mass.

The calculation string is straightforward, but I'll do the whole thing anyway because I need to show you the use of molarity as a conversion factor. Just like any conversion factor, it can be brought in right-side-up or upside-down.

OK, start with the given volume. Notice that it says 100.0 mL. Decimal shift that to liters, 0.1000 L, and start there.

L soln

Using molarity as a conversion factor, we convert this to moles of solute.

You will need  $\frac{2.19 \text{ mol NaCl}}{\text{L soln}}$  OR  $\frac{\text{L soln}}{2.19 \text{ mol NaCl}}$

You'll use the one on the left for this problem: that cancels L soln and gets you to mol NaCl.

L soln → mol NaCl

Molar mass then takes our moles of NaCl to grams. The molar mass for sodium chloride is 58.44 g.

L soln → mol NaCl → g NaCl

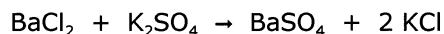
String it all out.

$$0.1000 \text{ L soln} \times \frac{2.19 \text{ mol NaCl}}{\text{L soln}} \times \frac{58.44 \text{ g NaCl}}{\text{mol NaCl}} = 12.8 \text{ g NaCl}$$

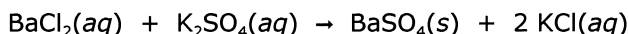
This tells us that we should dissolve 12.8 g NaCl in enough water until the solution volume is 100.0 mL. That gives us the required 2.19 M NaCl concentration.

### 9.3 Now we're ready for solution stoichiometry.

Let's jump right into a problem. For this we'll use the reaction of barium chloride with potassium sulfate to produce barium sulfate and potassium chloride.



Both reactants are soluble in water, so both of their solutions are homogeneous. One product, potassium chloride, is also soluble in water, but barium sulfate is not. Thus, the reaction produces a cloud of very fine particles of BaSO<sub>4</sub>, which eventually settle to the bottom of the pot. This is then a heterogeneous mixture. I can describe this with words, like I just did. Or I can say it with phases.



**Example 4.** You have 50.0 mL of a solution of K<sub>2</sub>SO<sub>4</sub> which has a concentration of 0.4963 M K<sub>2</sub>SO<sub>4</sub>. How many grams of barium chloride do you need to react with this amount of potassium sulfate?

The problem begins with the amount of one reactant and asks for an amount of a second reactant. This is the same kind of problem as we've done before. The problem uses the same four stoichiometry steps which we've been using. Step 1 is the balanced equation, which I provided this time. Stoichiometry Steps 2 - 4 are the same, except there will be a molarity involved.

Plot your path. You start with the given volume of the given reagent solution. It's in mL and you need L. Do your shift, 50.0 mL = 0.0500 L.

L soln

- Step 2. Convert this to moles, this time using molarity.

L soln → mol K<sub>2</sub>SO<sub>4</sub>

You will need  $\frac{0.4963 \text{ mol K}_2\text{SO}_4}{\text{L soln}}$  OR  $\frac{\text{L soln}}{0.4963 \text{ mol K}_2\text{SO}_4}$

- Step 3. Relate the moles of K<sub>2</sub>SO<sub>4</sub> to the moles of BaCl<sub>2</sub> by the rxn ratio (1:1).

L soln → mol K<sub>2</sub>SO<sub>4</sub> → mol BaCl<sub>2</sub>

- Step 4. Molar mass, molar mass, molar mass. For BaCl<sub>2</sub>, it's 208.2 g.

L soln → mol K<sub>2</sub>SO<sub>4</sub> → mol BaCl<sub>2</sub> → g BaCl<sub>2</sub>

- OK, string it.

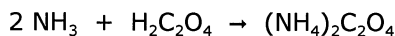
$$0.0500 \text{ L soln} \times \frac{0.4963 \text{ mol K}_2\text{SO}_4}{\text{L soln}} \times \frac{1 \text{ mol BaCl}_2}{1 \text{ mol K}_2\text{SO}_4} \times \frac{208.2 \text{ g BaCl}_2}{\text{mol BaCl}_2} = 5.17 \text{ g BaCl}_2$$

This says that the reaction needs 5.17 g barium chloride to react with the given amount of potassium sulfate solution.

Notice the similarity between the use of molarity and of molar mass in a stoichiometry string. Both relate a measurable amount to how many moles of formula units are present.

In our Example here, we used a volume of a solution of one reagent to find a mass of another reagent. The backwards can also happen. I'll show this with another example.

**Example 5.** The reaction of aqueous ammonia with oxalic acid, H<sub>2</sub>C<sub>2</sub>O<sub>4</sub>, produces the ionic compound ammonium oxalate according to the following balanced equation.



You need to make 24.63 g ammonium oxalate. You have plenty of oxalic acid which you will use as an excess reagent. You need to know how much ammonia solution to use as the limiting reagent. How many mL's of the ammonia solution do you need, if its concentration is 1.037 M NH<sub>3</sub>?

This is just standard stoichiometry. Step 1, the balanced equation, is provided. Start into the path, beginning with the indicated amount of ammonium oxalate to make.

g (NH<sub>4</sub>)<sub>2</sub>C<sub>2</sub>O<sub>4</sub>

- Step 2. Molar mass. For ammonium oxalate, it's 124.10 g.

g (NH<sub>4</sub>)<sub>2</sub>C<sub>2</sub>O<sub>4</sub> → mol (NH<sub>4</sub>)<sub>2</sub>C<sub>2</sub>O<sub>4</sub>

- Step 3. Rxn ratio, 2:1 or 1:2.

g (NH<sub>4</sub>)<sub>2</sub>C<sub>2</sub>O<sub>4</sub> → mol (NH<sub>4</sub>)<sub>2</sub>C<sub>2</sub>O<sub>4</sub> → mol NH<sub>3</sub>

- Step 4. Step 3 ended at mol NH<sub>3</sub>. The problem seeks volume of the ammonia solution. The connection is molarity.

g (NH<sub>4</sub>)<sub>2</sub>C<sub>2</sub>O<sub>4</sub> → mol (NH<sub>4</sub>)<sub>2</sub>C<sub>2</sub>O<sub>4</sub> → mol NH<sub>3</sub> → L soln

You will need  $\frac{1.037 \text{ mol NH}_3}{\text{L soln}}$  OR  $\frac{\text{L soln}}{1.037 \text{ mol NH}_3}$

Which one do you use? Think about it.

- Here we go.

$$\begin{array}{ccccccc}
 \text{g (NH}_4)_2\text{C}_2\text{O}_4 & \rightarrow & \text{mol (NH}_4)_2\text{C}_2\text{O}_4 & \rightarrow & \text{mol NH}_3 & \rightarrow & \text{L soln} \\
 24.63 \text{ g (NH}_4)_2\text{C}_2\text{O}_4 & \times & \frac{\text{mol (NH}_4)_2\text{C}_2\text{O}_4}{124.10 \text{ g (NH}_4)_2\text{C}_2\text{O}_4} & \times & \frac{2 \text{ mol NH}_3}{1 \text{ mol (NH}_4)_2\text{C}_2\text{O}_4} & \times & \frac{\text{L soln}}{1.037 \text{ mol NH}_3}
 \end{array}$$

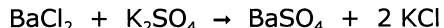
The number crunching comes to 0.3828 L soln, but don't forget that the problem asked for the answer in mL. Do your decimal shift. The answer is 382.8 mL: you need 382.8 mL of the aqueous ammonia solution to make the required amount of product.

Done.

By the way, everything in Chapter 8 is fair game with molarity in the problem. All of that stuff can still apply. In this last Example, I said we would run the reaction with oxalic acid as the excess reagent and we calculated ammonia to be the limiting reagent. I incorporated that terminology right into the problem. You can also do percent yield with molarity. Thus, in general, the use of a molarity does not change other stoichiometry aspects.

As a matter of fact, let's go ahead and do a percent yield problem.

**Example 6.** We'll use the  $\text{BaCl}_2/\text{K}_2\text{SO}_4$  reaction from earlier.



For a reaction using 50.0 mL of 0.4963 M  $\text{K}_2\text{SO}_4$  solution, 5.1 g  $\text{BaSO}_4$  were obtained after the work-up was completed. What is the percent yield?

As we discussed in Section 8.3, percent yield is actual over theoretical. We are told that the actual yield for this problem is 5.1 g. We need to calculate the theoretical yield, and that's a stoichiometry string. We need to find out how much  $\text{BaSO}_4$  can be made from the given amount of  $\text{K}_2\text{SO}_4$ . (By the way, this problem deals only with  $\text{K}_2\text{SO}_4$  and  $\text{BaSO}_4$ . Although earlier we had also worked with  $\text{BaCl}_2$ , that is not mentioned in the present problem and we leave it out here.)

Plot your path. Start with the given solution, but bring it in as L.

L soln

- Step 2. Convert to moles, using molarity.

L soln  $\rightarrow$  mol  $\text{K}_2\text{SO}_4$

- Step 3. Rxn ratio (1:1).

L soln  $\rightarrow$  mol  $\text{K}_2\text{SO}_4 \rightarrow$  mol  $\text{BaSO}_4$

- Step 4. Molar mass. For  $\text{BaSO}_4$ , it's 233.4 g.

L soln  $\rightarrow$  mol  $\text{K}_2\text{SO}_4 \rightarrow$  mol  $\text{BaSO}_4 \rightarrow$  g  $\text{BaSO}_4$

- OK, string it.

L soln  $\rightarrow$  mol  $\text{K}_2\text{SO}_4 \rightarrow$  mol  $\text{BaSO}_4 \rightarrow$  g  $\text{BaSO}_4$

$$0.0500 \text{ L soln} \times \text{—————} \times \text{—————} \times \text{—————} = \text{—————} \text{ g BaSO}_4$$

Enter your conversion factors and calculate the amount of  $\text{BaSO}_4$ . That's your theoretical yield. Plug that into the percent yield formula.

$$\begin{array}{c}
 \text{actual yield} \\
 \searrow \\
 \text{percent yield} = \frac{5.1 \text{ g}}{\text{g}} \times 100\% = \text{—————} \% \\
 \nearrow \\
 \text{theoretical yield}
 \end{array}$$

Your final answer should be 88%. Be sure you can get this. If not, find and fix the error.

#### 9.4 Recapitulate

This would be a very good time to recap. We have covered some very important turf.

We now have two ways of relating something which can be measured to the number of formula units in the sample. Molar mass connects a measurable mass with the number of units. Molarity connects a measurable volume with the number of units.

<p>Molar Mass</p> $\frac{\text{grams of compound}}{\text{moles of compound}}$	<p>Molarity</p> $\frac{\text{moles of compound}}{\text{liters of solution of that compound}}$
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Both can enter into stoichiometry strings as conversion factors, either right-side-up or upside-down, as needed. Notice that they differ in that molar mass has moles downstairs while molarity has moles upstairs. Molarity has moles upstairs because it was set up as a concentration term.

This ends our three-chapter sequence covering stoichiometry. It's going to take time to get comfortable with it. It's going to take practice, practice and more practice.

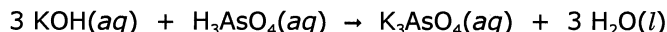
#### Problems

1. A solution contains 8.16 g NaHCO<sub>3</sub> in 200.0 mL. What is the molarity of the sodium bicarbonate?
2. What is the molarity of a solution which contains 15.7 g barium chloride in a volume of 100.0 mL?
3. You need to prepare 400. mL of 0.280 M K<sub>2</sub>CrO<sub>4</sub> solution. How many grams of K<sub>2</sub>CrO<sub>4</sub> are needed?
4. You need to prepare 150. mL of 2.7 M potassium nitrate. How many grams of potassium nitrate are needed?
5. The following equation is balanced.



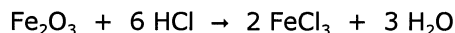
If the reaction is conducted using 18.4 mL of PCl<sub>3</sub> (density = 1.57 g/mL), how many grams of P<sub>4</sub>O<sub>6</sub> can be made?

6. The following equation is balanced.



What volume (in mL) of 0.550 M KOH is required in order to make 26.9 g K<sub>3</sub>AsO<sub>4</sub>?

7. The following equation is balanced.



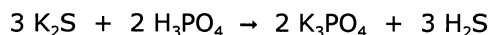
The reaction is conducted using 53.00 mL of 2.017 M HCl. How many grams of FeCl<sub>3</sub> can be produced?

8. The following equation is balanced.



How many grams of CaCO<sub>3</sub> are needed to react with 150.0 mL of 1.097 M HBr?

9. The following equation is balanced.



The reaction uses 147 mL of 2.019 M H<sub>3</sub>PO<sub>4</sub> solution. How many grams of K<sub>2</sub>S are needed?