Chapter 12: Stoichiometry – Fun to Say, Fun to Do

It's nearly impossible to come up with a good image to start a chapter about stoichiometry, so I figured I'd just include this picture of what appears to be an imprisoned orange tree.

http://commons.wikimedia.org/wiki/File:The_Mother_Orange_Tree.jpg

Chapter 12: Stoichiometry – Fun to Say, Fun to Do

You're probably reading this chapter because you're being forced to do so by an authority figure of some kind. Or perhaps you're reading it for the pure love and wonder of science. Most likely, you're reading it because some teacher keeps saying "stoy – key – ah – meh – tree" and you have no idea what he or she is talking about.

Well, wonder no further. We're going to talk about the magical world of stoichiometry and how it can be used to enrich your life, etc.

Section 12.1: What the Hell is Stoichiometry?

You probably don't have any clue what stoichiometry is, mainly because the word itself seems designed to make your brain melt right out of your ears. Fortunately, you have me, and I can give you this information in a way that's less likely to melt brains.

The word **stoichiometry** is just a fancy way of saying "the method you use to figure out how much of a chemical you can make, or how much you need, during a reaction." For example, if you're doing a reaction and want to make 88.5 grams of the product, you'd do a bunch of calculations to figure out how much of each reagent you'd need. Those calculations are stoichiometry.[1]

Irrelevant Information to Take Your Mind Off of Stoichiometry

Figure 12.1: *The first person to use the word "stoichiometry" was Nicephorus I, the ecumenical Patriarch of Constantinople in the early 9th century. However, the term as he used it refers to the number of lines of text in the New Testament. This suggests that the term "stoichiometry" has been annoying people for over a millennium.*

Image: Classical Numismatic Group, www.cngcoins.com

[1] They're actually referred to as "stoichiometric calculations", but since that doesn't exactly clear things up, I'll just call 'em stoichiometry.

As you've probably already discovered in your chemistry class, there are about a bazillion types of stoichiometric calculation out there. The good news, however, is that none of them are all that difficult. Seriously.

Section 12.2: Doing Stoichiometry

Before I show you some examples of stoichiometry, let me show you a handy picture that you'll be using a lot:[2]

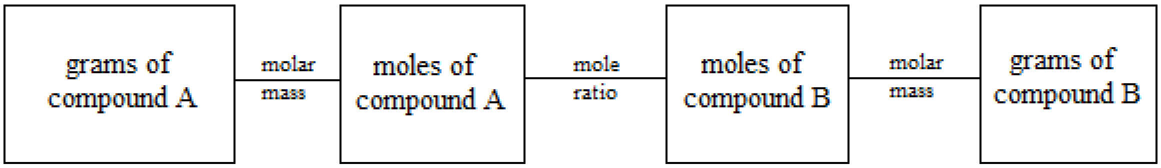


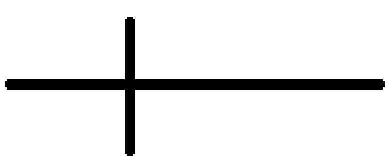
Figure 12.2: *A handy picture that you'll be using a lot.*

The best way to teach you how to use this picture to do stoichiometry is to simply give you a stoichiometry problem and solve it:

Problem: Given the equation $2\ H_2 + O_2 \rightarrow 2\ H_2O$, how many grams of water can be made with 50.0 grams of oxygen and an excess of hydrogen?[3]

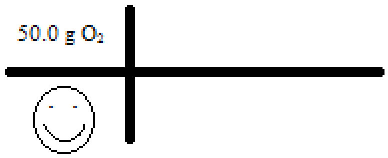
Answer: Let's go through the following steps to solve this problem:

1. **Determine what you're trying to figure out, and what you've been given:** In this problem, you've been given the value "50.0 grams of oxygen", so we'll call O_2 "compound A." The problem tells you that you're trying to find the "grams of water", so we'll call H_2O "compound B."[4]
2. **Figure out where you are on this table, and where you're trying to get:** Since we're given "50.0 grams of oxygen", we start in the "grams of compound A" box. Because we're trying to find out how many grams of water we'll be making, we end in the "grams of compound B" box.
3. **Make a "t":** You'll recognize the following steps from the mole calculation chapter (chapter 11).

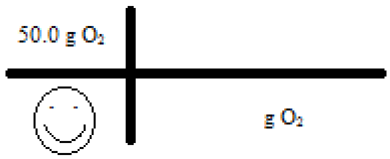


[2] That's my subtle way of saying "memorize this."
[3] When one of these problems refers to an "excess of [something]" what that really means is that there's so much of that compound hanging around that you really don't need to worry about it. Focus on the other reactant instead.
[4] It doesn't really matter which compound is A and which is B as long as you follow the steps in this guide.

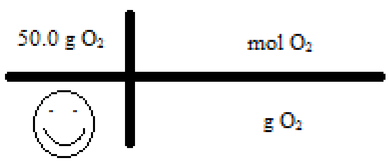
4. **Put the thing that you were given in the problem in the top left of the t.** Since you were given the value "50.0 grams of oxygen" in the problem, put that in the top left of the t:[5]



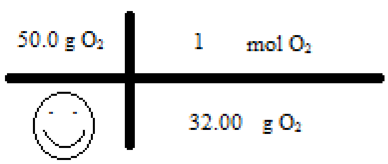
5. **Put the units of the thing that's in the top left in the space on the bottom right.** The units in the top left are "g O_2", so put "g O_2" in the bottom right:



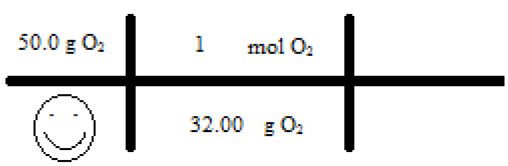
6. **Put the units of what you're trying to find in the top right.** Now, this one is a little more challenging. Obviously, we ultimately want to find the number of grams of water that can be formed. However, an examination of the table above tells us that we have to do several unit conversions to make that happen. As a result, our first conversion will simply be from grams of O_2 to moles of O_2:



7. **Put in the conversion factor between the things on the right.** The handy diagram says "molar mass", and molar mass is given to us in units of "grams in 1 mole", so we'll put a "1" in front of "mol O_2" and the molar mass of O_2 (32.00 g) in front of "g O_2":



8. Clearly, we're not done. In fact, all we've done is to do the same type of mole calculation that you learned about in Chapter 11. In order to get all the way to grams of H_2O, we'll have to do two more calculations. Let's extend the t-chart a little:



[5] The smiley face is put in the bottom left to make stoichiometry a happy and cheerful experience.

9. Move the units of the thing in the top left (mol O_2) to the bottom right. This is the same as step 5 above:

50.0 g O_2	1 mol O_2	
	32.00 g O_2	mol O_2

10. Put the units of what you're trying to find in the top right (same as step 6 above). Since we can't go directly to grams of water, we'll go to moles of water, as that's the next step on our handy chart:

50.0 g O_2	1 mol O_2	mol H_2O
	32.00 g O_2	mol O_2

11. Put in the conversion factors (same as step 7 above). In this step, the conversion factors are called the "mole ratio" because the conversion is from moles of one compound to moles of another compound. The conversion factors in the mole ratio step are the coefficients in front of each compound in the balanced equation:

50.0 g O_2	1 mol O_2	2 mol H_2O
	32.00 g O_2	1 mol O_2

12. Since we're not at grams of water yet, we need to go through the steps of extending the t-chart, writing the units from the top left in the bottom, and so forth (steps 4-7 again):

50.0 g O_2	1 mol O_2	2 mol H_2O	18.01 g H_2O
	32.00 g O_2	1 mol O_2	1 mol H_2O

13. In this, the last step, we multiply this stuff together like a series of fractions, where the stuff on the top is multiplied together and divided by the product of the stuff on the bottom multiplied together. This gives us an answer of 56.3 g H_2O.[6]

> ***Made-up Fun Fact***
>
> *Fractions were invented in 1875 by Patrick L. Fraction, a British accountant. Prior to that time, whenever somebody would cut a pie in two, they'd refer to each piece as "the product of a pie that hath been cleaved in twain." Because the many variations on that were annoying to say every time somebody took a piece of pie, Fraction came up with the convenient "fractions" that we use today.*

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[6] The significant figures in this calculation are determined by the value "50.0 g O_2". Though it might seem that the numbers of moles would give us only one significant figure, they're exact values (i.e. there are approximately 18.01 grams of water in *exactly* one mole of water), so we treat them as if they have infinite significant figures.

Section 12.3: Limiting Reagents[7]

In the example above, we assumed that we had a limited amount of oxygen and an excess amount of hydrogen to work with. However, when we do reactions in the real world, we usually don't have an unlimited amount of one of the reagents – we usually have limited amounts of each.

What this means is that when we do a reaction with two reagents, one of them usually runs out before the other. The reagent that runs out is called the **limiting reagent** because it's the one that puts a limit on how much of the product will be formed.

You can think of this conceptually by using a recipe for bread that my old alcoholic grandfather came up with. To make this bread, he'd add two cans of beer to one bag of crushed up potato chips and bake the resulting mixture in the oven to make a mangled, disturbing "loaf" of "bread".[8] Given this disturbing recipe, how many loaves of his bread could he make with 24 cans of beer and 11 bags of chips?

To solve this problem, most people will do the following:

- Figure out how much bread you can make from 24 cans of beer. (12 loaves)
- Figure out how much bread you can make from 11 bags of chips. (11 loaves)
- The smallest answer is the correct one – 11 loaves. The limiting reagent (the thing that you ran out of) is chips, and the excess reagent (the thing that didn't run out) is the beer.

Figure 12.2: *Delicious, delicious bread.*

http://commons.wikimedia.org/wiki/File:Essene_flat_Bread_100_pct_Wheat_Sproud.JPG

Likewise, let's say that we're going to perform the reaction $2\ H_2 + O_2 \rightarrow 2\ H_2O$. If I have 14.0 grams of H_2 and 48.0 grams of O_2, how many grams of water can I make? To find this out, I:

- Find out how many grams of water you can make from 14.0 grams of H_2.[9] (126 grams)
- Find out how many grams of water you can make from 48.0 grams of O_2. (54.0 grams)
- The smallest answer is the correct one – 54.0 grams. The limiting reagent is oxygen and the excess reagent is hydrogen.

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[7] The term "limiting reactant" is also sometimes used.
[8] I wish I was kidding about this.
[9] These calculations are done using stoichiometry as discussed in Section 12.2.

Excess Reagents

In the beer/chips/bread example, I mentioned that we ran out of chips and had leftover beer. The reagent in a chemical process that isn't completely used up is called the **excess reagent**. Likewise, in the water example above, we said that hydrogen is the excess reagent because the oxygen ran out first.

To figure out how much of the excess reagent is left over, use the following equation:

$$\begin{array}{c}\text{amount of excess}\\\text{reagent left over}\end{array} = \begin{array}{c}\text{original amount}\\\text{of excess reagent}\end{array} - \begin{array}{c}\text{original amount}\\\text{of excess reagent}\end{array}\left(\frac{\begin{array}{c}\text{amount of product}\\\text{that's actually formed}\end{array}}{\begin{array}{c}\text{amount of product that}\\\text{the excess reagent can make}\end{array}}\right)$$

Or, to put it another way:

$$\begin{array}{c}\text{amount of excess}\\\text{reagent left over}\end{array} = \begin{array}{c}\text{original amount}\\\text{of excess reagent}\end{array} - \begin{array}{c}\text{original amount}\\\text{of excess reagent}\end{array}\left(\frac{\begin{array}{c}\text{the small number from your}\\\text{stoichiometry calculations}\end{array}}{\begin{array}{c}\text{the big number from your}\\\text{stoichiometry calculations}\end{array}}\right)$$

In the water example, we found that we could make 126 grams of water using 14.0 grams of H_2 and 54.0 grams of water using 48.0 grams of O_2. Because H_2 is the excess reagent, we can find that the amount that's left over is equal to:

$$\begin{array}{c}\text{amount of excess}\\ H_2 \text{ left over}\end{array} = 14.0 \text{ grams} - 14.0 \text{ grams}\left(\frac{54.0 \text{ grams}}{126 \text{ grams}}\right)$$

$$= 8.0 \text{ grams } H_2$$

Water: Moist Friend of Mankind

Figure 12.3: *From the stupid drinks that hipsters pay $12 for (shown here) to the kiddie pools that the same hipsters fill in their living rooms to be "edgy", water has been both a friend and an enemy to mankind. I once saw on some TV show that the human body is mostly water. Well, it has a lot of water, anyway – something like 80%, but maybe less. I wasn't really listening, because Angry Birds had just come out and I thought that was way more interesting than watching a show about water. Seriously, like I'm going to watch an hour-long show about water.*

http://commons.wikimedia.org/wiki/File:Hot_chocolate_mug_with_whipped_cream.jpg

Section 12.4: Percent Yield

Let's say that, like everybody else in the world, you screw stuff up on a more-or-less continuous basis. This isn't to say that you're not a good person – it's just that you're human, and human beings screw up stuff all the time through either little or huge mistakes. If you're still under the impression that you're perfect, I've got bad news: Mom was lying to you when she said that you were her "perfect little angel."[10]

Given the presupposition that we all screw stuff up on a more-or-less continual basis, it stands to reason that we'll screw stuff up in the laboratory as well. One of the big questions we need to ask ourselves is "How did we manage to screw that up?" After all, if we know why we screwed something up we'll be less likely to screw it up in the future. The following types of error exist when performing chemical reactions:

- **Human error:** I put this first because it's the main cause of laboratory errors. Your friend Bobby might drop the crucible in the sink or you might accidentally think that you need 1.0 grams of a compound instead of 1.5 grams of a compound. The ways in which human beings can screw stuff up are varied and extremely creative. Human error can be partially accounted for by being extremely careful, but given that nobody is perfect, it will always exist to some extent or another.

Figure 12.4: *The 1979 partial meltdown of the Three Mile Island nuclear power plant was due to human error. Apparently, the operators couldn't figure out that there was a stuck valve, and eventually radioactive water was released into the environment. Reactor 1 wasn't affected, so it's still chugging away, making power.*

http://commons.wikimedia.org/wiki/File:Three_Mile_Island_(color).jpg

- **Systematic error:** These are errors that are made in the same way every time due to quirks or limitations in the procedure. For example, if you were to have a faulty stopwatch, you'll always let your reactions run for 30 seconds more than they should. If your lab partner spits in your beaker every time you start a reaction, your product will always contain his spit. You get the idea. This type of error can usually be traced back to some error in the procedure – in other words, it's human error as manifested in the instructions you give yourself.

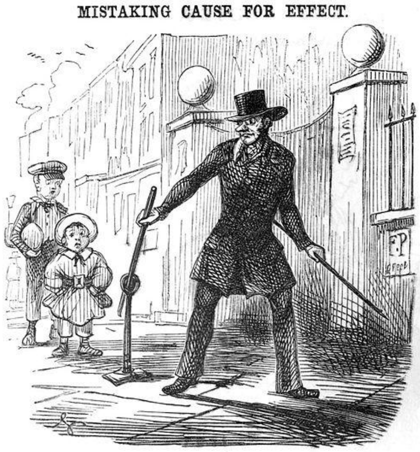
- **Instrumental/black box error:** You've got some machine and it gives you bad answers. These errors are called black box errors because the workings of the machines are mysterious and unknowable. For example, when you put a sample on a balance, the answer just shows up on a little computer screen. What happens inside of the balance? Who knows? It could be that the machine is humming along just fine, or it could be that the machine is full of rat poop and gives random answers. There's no fixing this type of error, though it can be minimized by calibrating

[10] More bad news: You're not actually an angel, either.

the machines ahead of time. Fortunately, this type of error is pretty uncommon with the stuff you'll see in a high school chemistry class, given the simplicity of most of the black boxes we use.

- **Unknown error:** You get done with the lab and you got a lousy answer. You were really careful with everything you did, so it wasn't human error. You made sure that your procedure was wonderful and that your equipment was all working properly. Well, here's the deal: Even though you think you did everything right, you still screwed up *somewhere*. Maybe Bobby still spits in the beaker and hasn't been telling you. Maybe the reagents were contaminated with pudding. Maybe somebody thought it would be funny to tape a penny to the bottom of your balance. All of this is my way of saying that sorry, bub – if you can't figure out what went wrong, it's probably a very well-hidden version of human error.

MISTAKING CAUSE FOR EFFECT.



Boy. "I say, Tommy, I'm blow'd if there isn't a Man a turning on the Cholera."

Figure 12.5: *The man in this cartoon doesn't understand how his actions are causing him to spread cholera. For that matter, neither do I, because the cartoon is from 1849 and doesn't really make any sense. You know how old cartoons are: There are just a bunch of people hanging around and one of them is doing something and everybody else looks surprised for no reason. Weird.*

http://commons.wikimedia.org/wiki/File:Mistaking_Cause_for_Effect_-_Turning_on_the_Cholera.jpg (public domain image)

Unfortunately, stoichiometric calculations assume that we're perfect in every way. If the calculation says that we'll make 75.0 grams of a compound, then we'll make 75.0 grams of a compound. What it doesn't take into account is the fact that there are inherent limitations to how well we can do things in the lab, and that these limitations result in smaller than expected quantities of product.

Fortunately, we have a way of figuring out how badly we screwed up: percent yield. The **percent yield** of a chemical reaction is a comparison of the amount of product you actually made, versus how much product that your stoichiometry calculations predicted you'd make. To find the percent yield of a reaction, use the following equation:

$$\text{percent yield} = \frac{\text{amount of product you actually made}}{\text{amount of product that stoichiometry predicted you'd make}} \times 100\%$$

What this means is that if your stoichiometry calculation predicted you'd make 75.0 grams of a compound and you actually made 58.0 grams of that compound in the lab, your percent yield would be:

$$\text{percent yield} = \frac{58.0 \text{ grams}}{75.0 \text{ grams}} \times 100\% = 77.3\%$$

Clearly, the better your percent yield, the fewer mistakes you made. For example, if you got a 10% yield, it means that 90% of your product is lost somewhere, never to be seen again. On the other hand, if you have a 90% yield, you've only lost 10% of your product, which is pretty good.[11] The only exception to this rule is if your answer is exactly 100.0% (which implies that you've made up your data because it's impossible to be perfect) or greater than 100% (which implies that you've either violated the law of conservation of mass or more likely that you have lots of impurities in your compound).

Awesome Errors in History

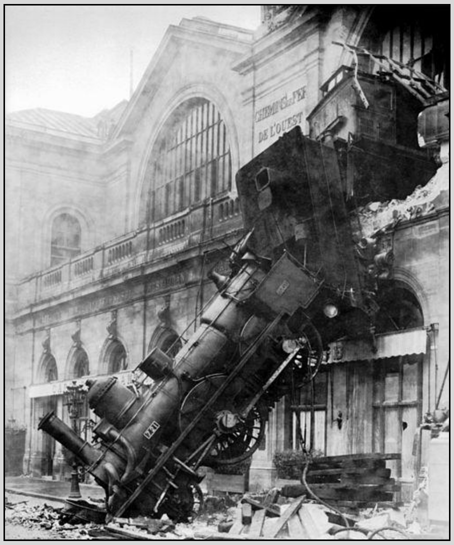


Figure 12.6: *This train ran off the end of the tracks at the Montparnasse rail station in 1895. Because the tracks were above ground, the train went out of the wall and onto the street below. The accident was caused by a defective brake. One person was killed: a woman who happened to be standing on the street below when the train crashed into her.*

Public domain image:
http://en.wikipedia.org/wiki/File:Train_wreck_at_Montparnasse_1895.jpg

[11] The actual relationship between percent yield and awesomeness is actually a lot more complicated than this. For example, some reactions don't ever reach completion, so it's *impossible* to get anywhere near 100% yield. Also, organic chemists frequently have multi-step reactions, where each step is likely to have about a 90% yield – those 90% terms really add up, making the final yield for the process very low even if each step was performed well. This is one of the reasons that many medications are so expensive – it takes a lot of chemical steps to make them, so you end up wasting a lot of stuff to get useful product.