

Chapter 15**AQUEOUS REACTIONS, Part 5**

©2004, 2008, 2011, 2022 Mark E. Noble

For the most part, we have completed our four major types of aqueous reactions: precipitation, neutralization, gas-forming and oxidation-reduction. I'm going to take some time now and combine these with stoichiometry. There's very little that I will do new in this Chapter. Yes, there will be a few new terms, but stoichiometry is still stoichiometry. Our four steps are the same as they were in Chapter 7.

- Step 1. Balance the equation.
- Step 2. Convert the given unit of the given reagent to moles.
- Step 3. Convert these moles of the given reagent to the moles of the desired reagent using the rxn ratio.
- Step 4. Convert the moles of the desired reagent to the desired units.

By this stage of the game, problems can get fairly complicated. One thing which you should look for is how to break the problem into parts. Although these problems can get ugly, they are just the sum of several parts. Look for the parts, and then do the parts step by step.

15.1 Examples

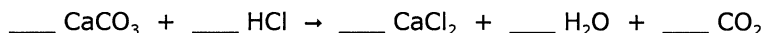
Let's get started.

Example 1. Since Chapter 11, we've talked a bit about stomach acid (hydrochloric acid) and I mentioned antacids which contain calcium carbonate. Let's put this together. What happens when that antacid hits your stomach? How many mL of 0.1053 M HCl will react with 1.000 g CaCO₃?

To do this, there are a number of things to consider.

- First of all, this is stoichiometry, since it's asking for a volume of something from a given amount of something else. In order to do stoichiometry, you need a balanced equation.
- In order to do the balanced equation, you need to know what's going on. So what is going on? Well, you have a carbonate and an acid, so that means this is a gas-forming reaction and carbon dioxide is the gas product.
- What else? From our coverage of gas-forming reactions, remember that a salt and water are also produced in this type of reaction. What salt? The anion comes from the acid, so the anion is chloride. The cation comes from the other reactant, so the cation is Ca²⁺. The salt is calcium chloride, CaCl₂.

We have enough to get started now. If you're lost already, go back to gas-forming reactions in Section 12.3. Otherwise, let's proceed to the balanced equation. I've lined everybody up; you fill in the coefficients.

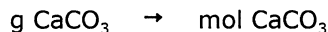


Now, on to number-crunching. You are given 1.000 g of CaCO₃; you are asked for the volume which is needed for the acid solution, knowing that the concentration is 0.1053 M HCl. We will set up the path for the stoichiometry string.

The given reagent is CaCO₃; it was provided as grams.



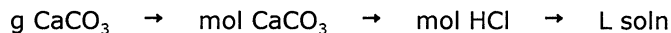
Take it to moles. That needs molar mass.



Go to moles of HCl. You need rxn ratio.



Finally, go to volume of the HCl solution using the known molarity.



Plug and chug.

$$\begin{array}{ccccccc} \text{g CaCO}_3 & \rightarrow & \text{mol CaCO}_3 & \rightarrow & \text{mol HCl} & \rightarrow & \text{L soln} \\ 1.000 \text{ g CaCO}_3 & \times & \frac{\text{mol CaCO}_3}{100.09 \text{ g CaCO}_3} & \times & \frac{2 \text{ mol HCl}}{\text{mol CaCO}_3} & \times & \frac{\text{L soln}}{0.1053 \text{ mol HCl}} = 0.1898 \text{ L soln} \end{array}$$

So there's your answer, almost. The problem wanted mL, not L: it's 189.8 mL. 1.000 g CaCO₃ can react with 189.8 mL of 0.1053 M HCl. Done.

Next.

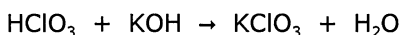
Example 2. You have 25.00 mL of 0.4062 M chloric acid. You react this with potassium hydroxide. How many grams of salt can be made by this process?

This one's quite tedious, too. Look for the parts.

- ▶ You're given amounts of one reagent and you're asked for amounts of something else. That's stoichiometry and that means you need a balanced equation.
- ▶ So what's going on here? Well, you have an acid, and its formula is HClO₃. You also have KOH, which is a base. That means this is a neutralization reaction.
- ▶ What are the products? In general, a salt and usually water. What salt? The cation comes from the base, so it's potassium ion. The anion comes from the acid, so it's chlorate. The salt is potassium chlorate, KClO₃.

Once again, see how we can break this problem into parts. That's a key feature to solving these more complicated types.

Now get started. Prepare the balanced equation from the formulas and from the information you have just gathered. Enter your reagents and balance things out.



The balance is easy in this case; everybody ends up with a coefficient of one.

Now construct the path. What is the given unit of what reagent? Well, you were told that you have 25.00 mL of chloric acid solution. You'll need that in liters, 0.02500 L.

L soln

Convert to moles. That connection is molarity.

L soln \rightarrow mol HClO₃

Convert those moles to moles of what you're looking for. You need the rxn ratio.

L soln \rightarrow mol HClO₃ \rightarrow mol KClO₃

And finally, convert those moles to grams of the salt. That needs molar mass.

$$\begin{array}{ccccccc} \text{L soln} & \rightarrow & \text{mol HClO}_3 & \rightarrow & \text{mol KClO}_3 & \rightarrow & \text{g KClO}_3 \\ 0.02500 \text{ L soln} & \times & \frac{0.4062 \text{ mol HClO}_3}{\text{L soln}} & \times & \frac{\text{mol KClO}_3}{\text{mol HClO}_3} & \times & \frac{122.55 \text{ g KClO}_3}{\text{mol KClO}_3} = 1.244 \text{ g KClO}_3 \end{array}$$

That's it. The reaction can produce 1.244 g KClO₃.

By the way, you've probably used potassium chlorate or you've seen it in use. It's used a lot in fireworks and matches where it's used as an oxidizing agent. For those applications, you want a very brisk flame initially; the oxygen in the air is not enough for this. So they add an oxidant such as KClO₃ for the needed effect.

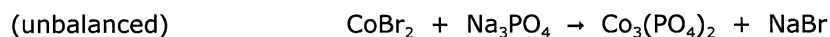
Again, if you're having problems with these problems, go back to the individual sections in the prior chapters. Otherwise we'll proceed.

Example 3. A solution containing 12.2 g cobalt(II) bromide was added to a solution of excess sodium phosphate. After work-up, 5.6 g of precipitate was obtained. What is the percent yield?

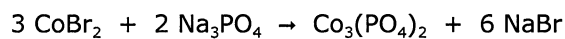
Here's another problem that might look like a disaster but which breaks down neatly into parts.

- ▶ Recall that percent yield is the actual yield over the theoretical yield. You are told that the actual yield is 5.6 g. Where do you get the theoretical yield? That comes from stoichiometry.
- ▶ What's the first thing you need? You need a balanced equation. You must derive it from the information given.
- ▶ In order to derive the balanced equation, you need some info about what's going on and you need formulas. Notice that you were told that a precipitate was obtained, so this is precipitation. The reactant formulas are CoBr_2 and Na_3PO_4 .
- ▶ What's the precipitate? Well, what ions are available? CoBr_2 will give Co^{2+} and Br^- ions; Na_3PO_4 will give Na^+ and PO_4^{3-} ions. From this mix of ions, what new combinations are possible? One possibility is NaBr and another possibility is $\text{Co}_3(\text{PO}_4)_2$. Soluble or insoluble? From the Solubility Trends, you can determine that NaBr is soluble and $\text{Co}_3(\text{PO}_4)_2$ is insoluble. Thus, $\text{Co}_3(\text{PO}_4)_2$ is the precipitate.

Now set up for a balanced equation. Line up the reactants and products.



Get your balance.



Proceed to the stoichiometry string. I'll just spell out the whole thing in one swell foop.

$$\begin{array}{ccccccc} \text{g CoBr}_2 & \rightarrow & \text{mol CoBr}_2 & \rightarrow & \text{mol Co}_3(\text{PO}_4)_2 & \rightarrow & \text{g Co}_3(\text{PO}_4)_2 \\ 12.2 \text{ g CoBr}_2 & \times & \frac{\text{mol CoBr}_2}{218.73 \text{ g CoBr}_2} & \times & \frac{\text{mol Co}_3(\text{PO}_4)_2}{3 \text{ mol CoBr}_2} & \times & \frac{366.73 \text{ g Co}_3(\text{PO}_4)_2}{\text{mol Co}_3(\text{PO}_4)_2} = \text{??? g Co}_3(\text{PO}_4)_2 \end{array}$$

Plug it in, punch it out, round it off. That's your theoretical yield; that's what you should get. The problem says that you only got 5.6 g in the end. Put these together to get percent yield.

$$\text{percent yield: } \frac{5.6 \text{ g}}{\text{??? g}} \times 100\% = 82\%$$

I gave the final answer for the percent yield, but the ??? is your part.

.....
Example 4. 26.65 g of sodium sulfate are to be prepared from the reaction of sulfuric acid with excess sodium hydroxide. How many grams of sulfuric acid are needed?

I'll let you sort through this one on your own. You'll need a balanced, undissociated equation.

(Does this reaction sound vaguely familiar? If you're having troubles, you might check Section 12.2.)

Plot your path.



You'll need some molar masses. Here are molar masses for all reagents in the balanced equation, in no particular order: 18.02, 40.00, 142.04 and 98.08. You figure out which ones you need.

Set up your string.

$$\underline{\hspace{2cm}} \times \underline{\hspace{2cm}} \times \underline{\hspace{2cm}} \times \underline{\hspace{2cm}} = \underline{\hspace{2cm}}$$

Plug and chug. I won't give you the answer directly, but here it is indirectly: without the decimal, the four sigfigs of the answer are the birth year of Monet and of Tchaikovsky.

This concludes our examples for now. Remember: on these big, more complicated problems, break the problem down. Look for steps. Think your way through them.

15.2 Titration

So far, there's nothing new in this Chapter. Now we head into some new stuff, although only the terminology is new. The aqueous chemistry is the same. The stoichiometry is the same. Only some words are new.

We're going to talk about titration. Titration is a very old and very common laboratory method of analysis. Many students in their first exposure to chemistry lab end up doing one or more titrations. I won't go into the details of the operation, since it is a laboratory procedure and it is better handled when you can see the actual glassware and the manipulations involved. I'm going to give a general outline, primarily to show some of the terminology and to show how this fits into where we're at right now.

A titration is conducted to measure the amount of something in a sample. Let's say you have some compound (let's just call it compound X) in some sample, and you want to know how much of that compound is in there. To do a titration, you add a reactant to the sample which will react with compound X; the added reactant is called a "titrant". This titrant is added as a solution and at a known concentration. You add the titrant solution slowly to your sample until stoichiometric amounts are present. The point at which stoichiometric amounts are present is called the "equivalence point". When you reach this point, you note how much volume of the titrant solution was used.

What does that get you? You know the titrant's concentration and you know the volume which was added to the sample. By stoichiometry, you can then calculate the amount of compound X which was present in the sample.

So why are we here? The most common titrations involve aqueous chemistry and the most common of these are acid-base and oxidation-reduction. Thus, titration is a fitting combination of aqueous reactions and of stoichiometry for our current coverage.

I'll do two examples of titration. The important thing to note for now is that a titration problem is primarily just another stoichiometry problem. There are some different terms involved, and there can also be aspects related to how a titration problem can be presented. I'll show you a bit of this. First, let me summarize the terms as we need them for here and now.

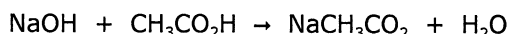
The verb is "titrate". It basically means to add something in portions and in a controlled way to react with something else. In a titration, you titrate your sample with your titrant. You continue to add titrant until you reach the equivalence point; a visual clue (called an "indicator") usually tells you when you reach that point. A related term is "end point", which is the same as equivalence point if done ideally. (There's a technical difference, but I'm not doing all the details here.) When you reach equivalence point (or end point), then you're done. This is the point of stoichiometric amounts and this is the point upon which you base your stoichiometry calculations.

Let's do a fairly typical titration problem: the analysis of acetic acid in vinegar.

Example 5. You have a solution which contains 4.00 g vinegar. You conduct an acid-base titration on this sample. The titrant is 0.1206 M NaOH, and it takes 24.51 mL of titrant to reach end point. How many grams of acetic acid were in the vinegar? What is the percent by mass of the acid in the vinegar?

Don't get thrown by the jargon. This is just stoichiometry.

First, set up your balanced equation. This is an acid-base reaction. Your reactants are $\text{CH}_3\text{CO}_2\text{H}$ and NaOH. You need a salt and usually water. The final balanced equation is



We now launch a standard stoichiometry string beginning with the known amount of the known reagent, which is the NaOH titrant. This string will get you the grams of acetic acid in the sample, which is the answer for the first question. Here's the string.

$$0.02451 \text{ L soln} \times \frac{0.1206 \text{ mol NaOH}}{\text{L soln}} \times \frac{\text{mol CH}_3\text{CO}_2\text{H}}{\text{mol NaOH}} \times \frac{60.05 \text{ g CH}_3\text{CO}_2\text{H}}{\text{mol CH}_3\text{CO}_2\text{H}}$$

Plug it in, punch it out, round it off. The answer is 0.1775 g: that's the mass of acetic acid which was present in the vinegar. Notice that we did not use the amount (4.00 g) of vinegar in this part: vinegar is itself a solution containing mostly water. The first question is only for the grams of acetic acid in that amount of vinegar.

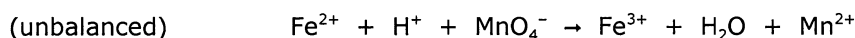
The second part is a little extra, since it asks for the percent by mass. I mentioned this back in Chapter 9 and you put a star there; go back and read that paragraph. Percent by mass requires the total mass of the sample, which means the total mass of all components in the sample. This is where you need the 4.00 g amount for the vinegar sample.

$$\frac{\text{mass of CH}_3\text{CO}_2\text{H}}{\text{masses of all components}} \times 100\% = \frac{0.1775 \text{ g}}{4.00 \text{ g}} \times 100\% = 4.44\%$$

This calculation tells us that acetic acid is 4.44% by mass in that sample of vinegar. That ends this Example.

Let's do another titration calculation. We'll do redox this time. An oldie-but-goodie is the titration of iron(II) ion using permanganate.

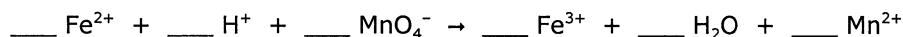
Example 6. You have a solution which contains some iron(II) sulfate and you want to know how much is in there. You titrate a sample of the solution, using 30.22 mL of 0.02032 M MnO_4^- solution to reach equivalence point. How many grams of FeSO_4 are in the sample? The unbalanced equation is the following.



Here's the setup for the problem.

This is again stoichiometry, but there's a slight twist to this problem. You'll start with the volume of titrant (MnO_4^- solution) and its molarity, but you can't use rxn ratio to get to mol FeSO_4 directly, because FeSO_4 is not in the balanced equation. The equation is in the net ionic form using Fe^{2+} , so rxn ratio will get you mol Fe^{2+} . (Where's the sulfate in the equation? Since sulfate is not shown in the equation, then it must be spectating.) We take FeSO_4 to be dissociated into Fe^{2+} ions and SO_4^{2-} ions, so that means every mole of Fe^{2+} came from the dissolved FeSO_4 . This will connect us to g FeSO_4 , which is the requested information. This may sound confusing, but watch how this comes out below.

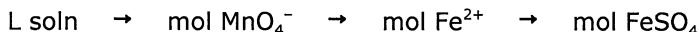
Alright, balance the equation. You work through it. I'll leave space. I'll also give a clue: you can get a good idea of the redox atoms from the wording of the setup.



Now get your stoichiometry string going. It will start out the usual way from the given amount of the titrant, which is the permanganate solution. From there, concentration gets you mol MnO_4^- and rxn ratio gets you mol Fe^{2+} .



At this time, we connect mol Fe^{2+} to mol FeSO_4 : the compound's formula says that there's one mole of Fe^{2+} ions per mole of FeSO_4 formula units.



Then you can go to grams FeSO_4 using molar mass.



Now go for it.

$$0.03022 \text{ L soln} \times \frac{0.02032 \text{ mol MnO}_4^-}{\text{L soln}} \times \frac{??? \text{ mol Fe}^{2+}}{??? \text{ mol MnO}_4^-} \times \frac{\text{mol FeSO}_4}{\text{mol Fe}^{2+}} \times \frac{151.91 \text{ g FeSO}_4}{\text{mol FeSO}_4}$$

You fix the rxn ratio. It comes from the balanced equation.

Plug it all in, punch it out and round it off. That's your answer. That's the number of grams of iron(II) sulfate originally present in the sample.

_____ g FeSO_4

I'm not giving you the answer, but I'll give you a different clue: ignoring the decimal, the four significant figures of the answer are a numeric palindrome.

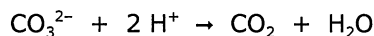
This ends our discussion of stoichiometry and of aqueous chemistry.

We've covered some fairly hefty problems in this Chapter. These are meant to be samplers. There are numerous variations to this, and your instructor may cover these. Again, some of these can get fairly complicated. Look at the problem and try to break it into parts. If a balanced equation is not given, then figure out how you can get one, since you'll probably need it. You need to be able to stare these down, no matter how menacing they appear. It doesn't matter if the problem looks like something from outer space; it doesn't matter if the face of doom is staring you straight in the eyeballs; it doesn't matter if it looks like the end of the world. You can do these. Just remember those saving words.

Klaatu. Barada. Nikto.

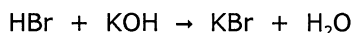
Problems

250. mL of 0.223 M $K_2C_2O_4$ are added to an aqueous solution of excess $CaCl_2$. How many grams of precipitate can be formed?
- 12.6 g KOH are reacted with excess HNO_3 . How many grams of salt can be formed?
- 25.0 mL of a solution of 1.04 M $Ba(OH)_2$ are added to 40.0 mL of a solution of 1.64 M $HClO_3$. How many grams of salt can be produced?
- 21.0 g of magnesium acetate were dissolved in water. This solution was added to a solution of excess sodium carbonate. After all work-up steps were completed, the actual yield of precipitate was 7.9 g. What was the percent yield?
- An aqueous solution of sodium carbonate was titrated using hydrochloric acid, according to the following net ionic equation.



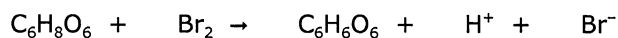
The titration required 21.39 mL of 0.1472 M hydrochloric acid to reach equivalence point. How many grams of sodium carbonate were present in the sample?

- 8.00 g of hydrobromic acid solution was titrated using a solution of potassium hydroxide, according to the following equation.



The potassium hydroxide concentration was 0.2316 M, and the titration required 28.70 mL to reach end point. How many grams of HBr were present in the sample? What was the percent by mass of HBr in the sample?

- Vitamin C (ascorbic acid), $C_6H_8O_6$, is an important antioxidant for humans. (An "antioxidant" is a reductant.) This vitamin can be bought in stores as a nutritional supplement, typically as tablets or as a powder. Upon prolonged standing, vitamin C can react with air, causing a decrease in purity. A sample of 22.8 mg of vitamin C powder was titrated in order to determine its purity, according to the following, unbalanced, redox equation.



The titration required 18.60 mL of 0.00617 M $Br_2(aq)$ solution to reach end point. How many milligrams of vitamin C were contained in the sample? What is the percent by mass of vitamin C in the sample?