Chapter 9: Solutions and Solution Reactions.

When the exact composition of a mixture is known, it is called a ***solution.*** In principle, a solution can be a solid, a liquid, or a gas. However, in the laboratory you commonly work with water (aqueous) solutions, and we will focus on aqueous solutions for the rest of this chapter.

Why do chemists dissolve substances in water? There are two reasons for using solutions, instead of pure substances:

1. Many reactions release a great deal of heat. The water absorbs this heat, and moderates the temperature of the reaction.

2. The reaction speed (the *reaction rate*) depends on the phase (solid, liquid, gas) of the reactants. Consider the following general reaction:

$$AB + C \rightarrow AC + B$$

Imagine the two reactants are solids (Figure 9.1). Each particle of the two reactants contains many individual atoms or molecules of "AB" or "C". Most of the molecules/atoms are inside the particle; the only places where AB and C can react with each other are the surfaces where they are touching. The vast majority of AB and C particles are not reacting at all. As a rule, reactions involving solids tend to be the slowest. It isn't unusual to find solid phase reactions taking years, centuries, or millennia to complete.

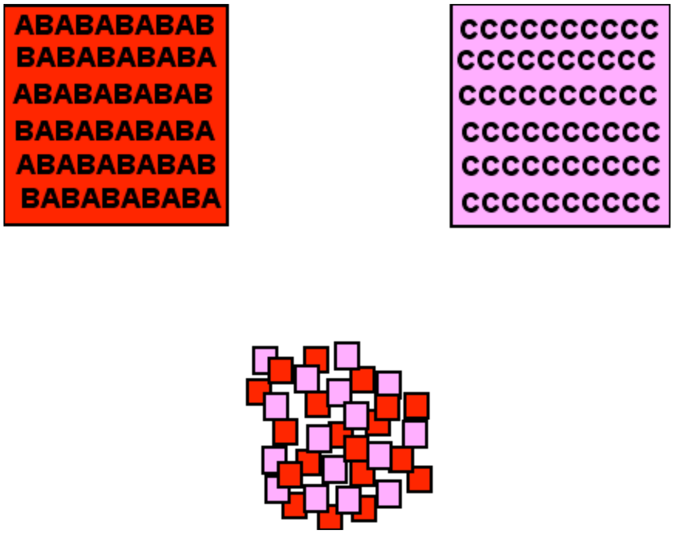

Figure 9.1. The reaction of solid particles of AB with C.

Now imagine the two reactants are gases. In a gas, all atoms or molecules exist as individual particles, and the individual particles mix easily and completely (Figure 9.2). No atom or molecule is trapped inside of larger particles, and all atoms or molecules can react. Gas phase reactions are generally the fastest.

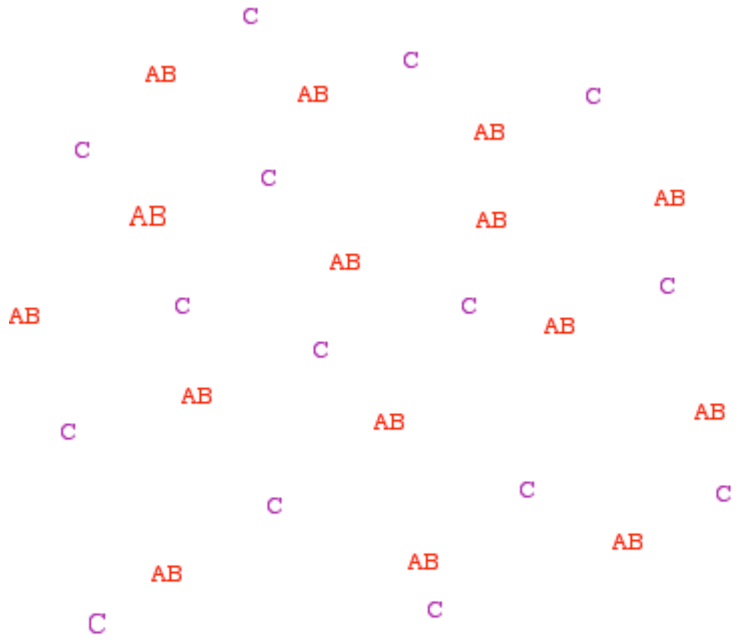

Figure 9.2. Reaction of AB and C in the gas phase.

So, if we want the fastest reactions possible, we should heat all of our reactants until they become gases! Unfortunately, the real world is not that simple, and the general procedure of heating substances until they become gases runs into two significant problems:

1. The boiling point of many substances is extremely high. For example, the boiling point of iron is 2750 °C, while the boiling point of sodium chloride is 1413 °C. Providing sufficient energy to boil these substances is a considerable problem, and handling them safely at these high temperatures is not simple.

2. Many substances decompose, or react with oxygen before reaching their boiling point. Ordinary sugar, for example, does not boil at atmospheric pressure – it decomposes into a brown mixture of melted sugar and carbon (it ***caramelizes***).

Liquid solutions give us most of the benefits of gases, without the need to vaporize the reactants. Nevertheless, using solutions instead of pure substances introduces complications we must address. The first, most obvious, complication is that we don't have one single pure substance – instead we have a mixture of substances.

Before going any farther, we need to introduce some new vocabulary terms.

All solutions are composed of a ***solvent***, and one or more ***solutes***. The solvent is the substance with the greatest mass percentage. The solutes are all other substances present in the solution – anything that is not the solvent. Consider the following solutions (Table 9.1):

Solution "A"	Solution "B"	Solution "C"	Solution "D"
1.00% sugar	25.00% sugar	55.00% sugar	78% nitrogen (N_2)
99.00% water	12.00% salt	45.00% water	21% oxygen (O_2)
	12.00% dye		1% trace (CO_2, Ar, Etc.)
	8.00% flavoring		
	43.00% water		

Table 9.1. Different solutions.

In solutions "A" and "B", water is the solvent; it is present in the greatest percentage. In solution "C", sugar is the solvent, and it is proper to describe solution C as "a 45% solution of water dissolved in sugar", although almost nobody does so. Solution "D" is a gas mixture, which is a special case. Sometimes two substances mix completely and readily, regardless of their proportions. For example, any amount of nitrogen gas mixes with any amount of oxygen gas. These substances are said to be ***miscible***, and we generally don't designate a solvent or solutes for miscible mixtures.

Any combination of solvent and solute produces a solution, but not all solutions are equivalent. A given solvent (i.e. water) dissolves a finite amount of each solute at fixed temperatures and pressures. The maximum solubility of sodium chloride in water is 359 g/L at 25 ^oC, while the maximum solubility of sugar (sucrose) in water at 25 ^oC is 2000 g/L. As a general rule, solubility increases with increasing temperature. However, there are exceptions to this rule, and most gases have lower water solubilities at higher temperatures.

If a solution contains less than the maximum amount of solute, then the solution is called ***unsaturated.*** If a solution contains the maximum amount of solute, then the solution is called ***saturated.*** A solution containing more than the maximum amount of solute is called ***supersaturated***.

"Ha Ha!," you say, "I have finally caught you! How exactly do you have more than the maximum amount of anything?"

"Easily," I say, "when you start with a hot solution and let it cool."

As I said before, solubility generally increases with increasing temperature. Figure 9.3 shows the effect of increasing temperature on the solubility of sodium acetate in water. The sodium acetate solubility is about three times greater at 100 ^{o}C than at 25 ^{o}C. Imagine that we make saturated sodium acetate at 100 ^{o}C, and allow it to slowly cool – what happens?

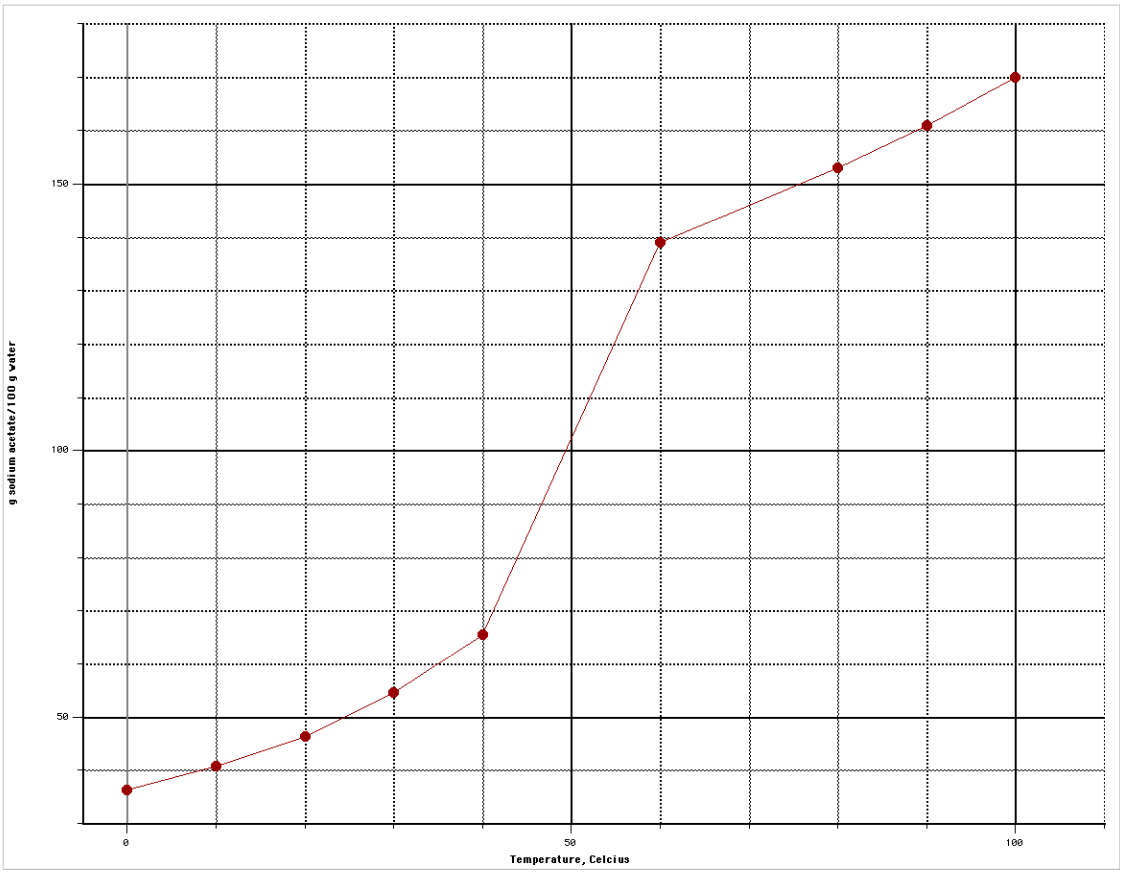

Figure 9.3. Solubility of sodium acetate (g/100 g water) versus temperature (^{o}C).

One possibility is that, as the solution cools, some sodium acetate ***precipitates*** from solution. Precipitation occurs whenever a solid material deposits from a liquid solution. When the solution has cooled to 25 ^{o}C, we have a saturated solution of sodium acetate, with a large amount of solid sodium acetate at the bottom of the container. The solid substance at the bottom of a saturated solution is called the ***precipitate***. This is the "normal" situation, and the general method used to make most saturated solutions.

But sometimes, as the solution cools, none of the sodium acetate precipitates. When this happens, the solution is supersaturated. Supersaturation is relatively rare, difficult to achieve, and extremely unstable! Almost any disturbance (a dust particle, scratching the inside of the container with a stirring rod, even loud noises) will cause supersaturated sodium acetate to precipitate. When precipitation occurs, we end up with the "normal" situation – a saturated solution of sodium acetate with a large amount of solid sodium acetate precipitate at the bottom of the container.

Solubility of ionic compounds.

It is often convenient for us to know whether a particular substance is soluble in water or if it is insoluble. First, we must make a distinction between these two terms. Generally, if a solute dissolves at least 1%, it is considered ***soluble***. Solubility depends not only on the solute but also the solvent: salt is soluble in water, but not very soluble in hexane (a common, nonpolar, organic solvent). If the solubility of a solute is <1%, we consider it ***insoluble***.

For ionic compounds, there is a relatively simple set of solubility rules that allow us to predict whether or not an ionic compound will be soluble in water, just from the chemical formula of the compound.

Solubility rules:

1. All group 1A metal (Li, Na, K, Rb, Cs, Fr) or ammonium (NH_4^+) salts are soluble. (As a practical matter, you will never encounter francium compounds, and rarely encounter rubidium or cesium compounds.)

2. All nitrate (NO_3^-), chlorate (ClO_3^-), perchlorate (ClO_4^-), or acetate ($CH_3CO_2^-$) salts are soluble. Nitrate and acetate are the two most important; chlorate and perchlorate are relatively rare.

3. Chloride, bromide, or iodide (Cl^-, Br^-, I^-) salts are soluble. There are three common exceptions: silver (Ag^+), mercury (I) (Hg_2^{+2}), and lead (Pb^{+2}) salts of these anions are insoluble.

4. Sulfate salts (SO_4^{-2}) are generally soluble. There are five common exceptions; Sr^{+2}, Ba^{+2}, Hg_2^{+2}, Pb^{+2} and Ca^{+2}: these salts are insoluble.

5. Hydroxide (OH^-) salts are generally insoluble. Exceptions are group 1A metal salts (see 1 above). (Sometimes Ca^{+2}, Ba^{+2}, and Sr^{+2} hydroxides are considered soluble.)

6. Sulfide (S^{-2}) salts are insoluble with the following exceptions: NH_4^+, group 1A metals. (Ca^{+2}, Sr^{+2} and Ba^{+2} are sometimes included.)

7. Carbonate (CO_3^{-2}), sulfite (SO_3^{-2}), chromate (CrO_4^{-2}), and phosphate (PO_4^{-3}) salts are insoluble with the exception of NH_4^+ and group 1A metals.

To use these rules, first look at the formula of the ionic compound and identify which ions are present. Then compare the ions present with the solubility rules, in the order presented, and note whether or not one of the ions present in the compound corresponds with one of the solubility the rules. You do not need to compare all ions with all rules: the first rule that tells you the compound is soluble or insoluble is all you need to consider.

For example, consider sodium sulfate (Na_2SO_4). This salt is composed of sodium ion (Na^+) and sulfate ion (SO_4^{-2}). Rule 1 indicates that this salt is soluble based upon the presence of sodium ion. (Rule 4 also indicates that this salt is soluble based upon the presence of sulfate ion and the absence of any exceptions, but is irrelevant – rule 1 takes precedence).

Iron (III) nitrate, $Fe(NO_3)_2$ is also soluble in water. There is no specific rule for iron compounds, but all nitrate compounds are soluble (Rule 2).

Nickel hydroxide, $Ni(OH)_2$ is insoluble by rule 5.

Sodium hydroxide is soluble by rule 1.

Concentration.

Whatever our solutions may be, we need some way to accurately and clearly express the amount of solute(s) dissolved in a given amount of solution. ***Concentration*** expresses the amount of a given solute present in a specified amount of a solution. Many different concentration units are used, but some units are more useful than others. One common unit is percentage – the ratio of the amount of solute present in 100 parts of solution. There are three different percentages that can be used, and they are not directly interchangeable.

Percentage by mass (**% m/m**) indicates the mass of solute present in 100 grams of solution. This is a convenient unit to use, and makes preparing solutions relatively simple. For example, if I want a 5.00% m/m solution of sodium chloride in water, I can combine 5.00 grams of sodium chloride and 95.00 grams of water. Percentage by mass takes advantage of the fact that mass always adds directly: 1.00 gram + 1.00 gram = 2.00 grams.

Preparing solutions whose concentrations are in % m/m is extremely easy. Simply weigh out masses of the solutes in grams equal to their percentage, and dissolve in enough water to make 100 grams of solution. For example, imagine I want a solution to have 1.5% sodium chloride, 1.7 % sodium sulfate, and 5.5% glucose ($C_6H_{12}O_6$) in water. How do I make it? Since "percent" means per 100, I can start with the idea of making 100 grams of solution. I need 1.5 grams sodium chloride, 1.7 grams sodium sulfate, and 5.5 grams glucose. This is a total of 8.7 grams, so I need 91.3 grams water. I add the solids solutes to the water and stir – my solution is made.

Percentage by volume **(% v/v)** indicates the volume of solute present in 100 mL of solution. This is a convenient percentage unit to use when your solutes are liquids, but there are some inherent problems. Volumes of liquids do not necessarily add directly: in some cases 10.00 mL + 10.00 mL is not 20.00 mL. This is frequently seen with mixtures of alcohol and water. If you have a 40% v/v solution of ethyl alcohol in water, it will be about 31.6% m/m. A 40% m/m solution of ethyl alcohol in water produces a 50.7% v/v solution.

Why don't volumes add directly, just like masses? Molecules have finite volume, and there are empty spaces between the molecules. Consider Figure 9.4, showing a methanol molecule (CH_3OH). This molecule has a definite physical shape and occupies space. When several methanol molecules are packed closely together, the packing isn't 100% efficient – some empty space is available between molecules. Any smaller molecule can slip into this empty space. When alcohol is mixed with water, the smaller water molecules can fit into empty space between alcohol molecules, using this otherwise wasted empty space.

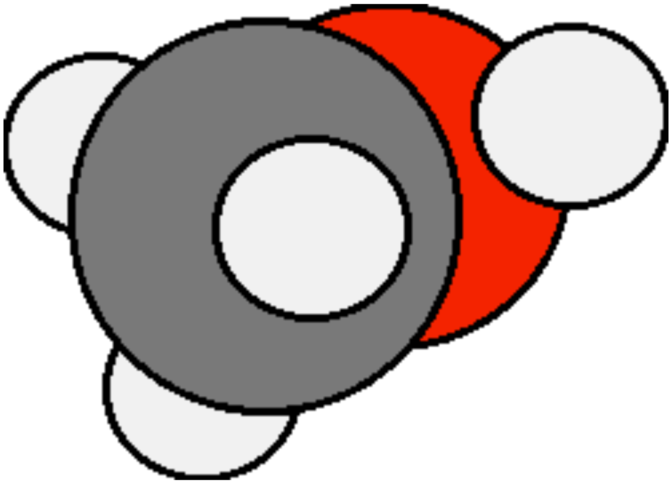

Figure 9.4. Methanol molecule. Carbon is dark gray, hydrogen light gray, oxygen red.

For a simple, "real world" example of this kind of packing, fill a glass with marbles (or beans, or any reasonably large objects) and then pour salt into the glass. The salt represents smaller molecules packing around larger molecules.

In both cases, it is possible to use the empty space between larger objects. As a consequence of this empty space between molecules:

$$10.00 \text{ mL large molecules} + 10.00 \text{ mL small molecules} \neq 20.00 \text{ mL total molecules}$$

Preparing % v/v solutions is also fairly simple. Imagine I wanted to make 500.0 ml of 15% v/v ethanol in water. The volume of ethanol I need is:

$$Volume = \left(\frac{15.00\%}{100}\right) \times 500.0mL = 75.00mL$$

I measure out 75.00 mL of ethanol, put it in a container and add enough water to make 500.0 mL of total solution. This will be *close* to 425 mL of water, but *not exactly* 425 mL.

Percentage by mass-volume (**% m/v**) is the mass of solute present in 100 mL of solution. A solution having 5.00 grams of sugar in 100 mL of solution is 5% m/v, or it would be if it were proper to report the percentage that way. There are two special things you need to know about this unit. First, percentages are not reported as just "%", but as % g/mL. Second, it is entirely possible to have mass-volume percentages greater than 100% w/v. For example, an 80% m/m solution of sulfuric acid in water contains 80 grams of sulfuric acid and 20 grams of water (sulfuric acid and water are miscible). The density of this solution is 1.73 grams/mL, so 100 grams of solution has a volume of 57.8 mL. The mass-volume percentage of sulfuric acid is therefore (80 grams/57.8 mL) x 100 = 138% m/v, which is confusing and ridiculous.

I strongly encourage you NOT to use mass-volume percentage. It is just as easy to express concentrations as grams/mL (or with appropriate conversions, other metric ratios such as mg/L, etc.). A solution containing 5.00 grams of sugar in 100 mL of solution has a sugar concentration of 0.05 g/mL, while 80 grams of sulfuric acid in 57.8 mL of solution has a concentration of 1.38 g/mL. There is nothing "magic" about percentages, and in this case you should avoid using a m/v percentage because it is confusing.

Regardless of whether you take my advice or not, I hope that you see that simply saying "The solution contains 3.00% "A" dissolved in "B"" is not enough information – we need to know which specific percentage you are describing.

However, if after all my encouragement you still want to use this silly concentration value then I will give you the procedure. Imagine you want to make 500.0 mL of a 5.0 % m/v solution of sodium chloride. How do you do it?

First, you need to weigh out some sodium chloride. The mass of sodium chloride you need is:

$$Mass = \left(\frac{5.0\%}{100}\right) \times 500.0mL = 25g$$

You put the 25g sodium chloride in an appropriate container, dissolve it in about 100 mL of water, and then add additional water until you have 500.0 mL total solution volume.

Molarity.

In Chapter 7, we learned about the mole, and how to convert the mass of a substance into the corresponding number of moles of substance. We can take this idea one step further. Chemists have defined a concentration measurement called ***molarity***, abbreviated as **M**, and having units of ***moles*/L**. The volume must be in liters, and it is the total volume of solution – not just the volume of water used to dissolve the substance! This is probably the most generally used concentration unit in chemistry, and it has a very specific advantage. When a given volume of solution is used, a specific number of moles of solute have been measured out.

Just as a chemist counts out atoms and molecules by weighing substances on a balance, the same chemist dispenses a known number of atoms/molecules from a solution by knowing the molarity of the solution. Let's look at a simple example. Consider a solution made by dissolving 90 grams of sugar ($C_6H_{12}O_6$) in enough water to make 1.00 L of solution.

Can we express the concentration in % m/m? Not very easily – while we know the mass of sugar, we don't know the mass of water used. I'm pretty sure we won't use 1000 grams of water, but that isn't much help. Can we express the concentration in % v/v? We know we are making 1.00 L of solution, but we don't know the volume of 90 grams of sugar. Besides, sugar is a solid, and % v/v is most useful when the substances are all liquids. We could use mass-volume percentage (if you insisted), although it would be better to label the solution as 90 g/L sugar (or 0.090 g/mL sugar).

Or we could calculate the molarity of the sugar solution. The molecular weight of sugar is 180.158 g/mole, and therefore 90 g sugar is 0.50 moles. The solution concentration is 0.50-moles/1.0L or 0.50 M sugar.

When I take fixed volumes of this sugar solution, each volume contains a fixed number of sugar molecules. I can calculate the number of moles of sugar present in each volume by multiplying the volume portion (in liters) by the molarity (Table 9.2).

1000 mL of solution (1.00 L) =	0.50 moles of sugar molecules
500 mL of solution (0.500 L) =	0.25 moles of sugar molecules
100 mL of solution (0.100 L) =	0.050 moles of sugar molecules
10 mL of solution (0.010 L) =	0.0050 moles of sugar molecules

Table 9.2. Volumes of solution and corresponding numbers of moles, for a 0.50M sugar in water solution.

$$\text{molarity} = \frac{\text{moles}}{\text{liters}}$$

$$M = \frac{m}{V}$$

Some solutions are prepared by diluting a fixed volume of a starting solution to a final volume, using water or some other liquid. Calculating the final molarity of the diluted solution is easy using the formula:

$$M_i V_i = M_f V_f$$

Where M_i is the initial concentration, V_i is the volume of starting solution, M_f is the final concentration, and V_f is the final volume of the diluted solution. NOTE: the concentrations must be molarity, or some other unit specifically using total volume of solution. Mass percent won't work, but volume percent and mass-volume percent will work just fine.

Molarity is the most commonly encountered concentration unit, and you must learn this unit. If you have to memorize this unit by brute force, then that is OK; sit in front of a mirror and chant "Molarity is ***moles*** per liter of solution" 500 times. If you have to go to a tattoo parlor and have the words "Molarity is ***moles*** per liter of solution" tattooed on your arm, then that is OK. Never forget that "Molarity is ***moles*** per liter of solution".

Solution stoichiometry.

In chapter 7, you learned to perform stoichiometry calculations using balanced chemical equations. Solution stoichiometry problems are similar, but slightly more complex because a solution contains the reactant, plus the solvent. In a sodium chloride solution, sodium chloride will probably be the reactant, while the water solvent is not chemically reacting. Simply measuring out 50.00 grams of solution doesn't mean that all 50.00 grams will react.

One simple way of handling this problem is to use the molarity and volume of reacting solutions to calculate the moles of reactants. Once we have moles, the ordinary stoichiometry calculations apply. For example, consider the reaction given below:

$$CaCl_2 + 2AgNO_3 \rightarrow 2AgCl + Ca(NO_3)_2$$

Imagine I have 0.025M $CaCl_2$, and 0.015M $AgNO_3$. If I use 50.00 mL of calcium chloride solution, what volume of silver nitrate do I need to completely react?

First, calculate the moles of calcium chloride available:

$$Molarity = \frac{moles}{Liters(solution)}$$

$$Molarity \times Liters(solution) = moles$$

$$0.025M \times 0.05000L = 0.00125moles\ \mathrm{CaCl_2}$$

Second, use the ratio of coefficients in the balanced chemical equation to calculate the moles of silver nitrate required:

$$0.00125moles\ \mathrm{CaCl_2} \times \frac{2\ \mathrm{AgNO_3}}{1\ \mathrm{CaCl_2}} = 0.0025moles\ \mathrm{AgNO_3}$$

Last, using moles of silver nitrate and concentration of silver nitrate, calculate the volume of silver nitrate needed to give the required number of moles:

$$Molarity = \frac{moles}{Liters(solution)}$$

$$0.015M\ \mathrm{AgNO_3} = \frac{0.0025moles\ \mathrm{AgNO_3}}{L}$$

$$L = \frac{0.0025moles\ \mathrm{AgNO_3}}{0.015M\ \mathrm{AgNO_3}} = 0.1667L = 166.7mL$$

(There are short cuts that can be used, but new students in chemistry don't need the short cuts. Learn to calculate properly using the long way – the short cuts can come later when you have more experience with chemical equations.)

Solubility mechanisms.

Solid ionic compounds generally have a uniform, repeating structure as shown in Figure 9.5.

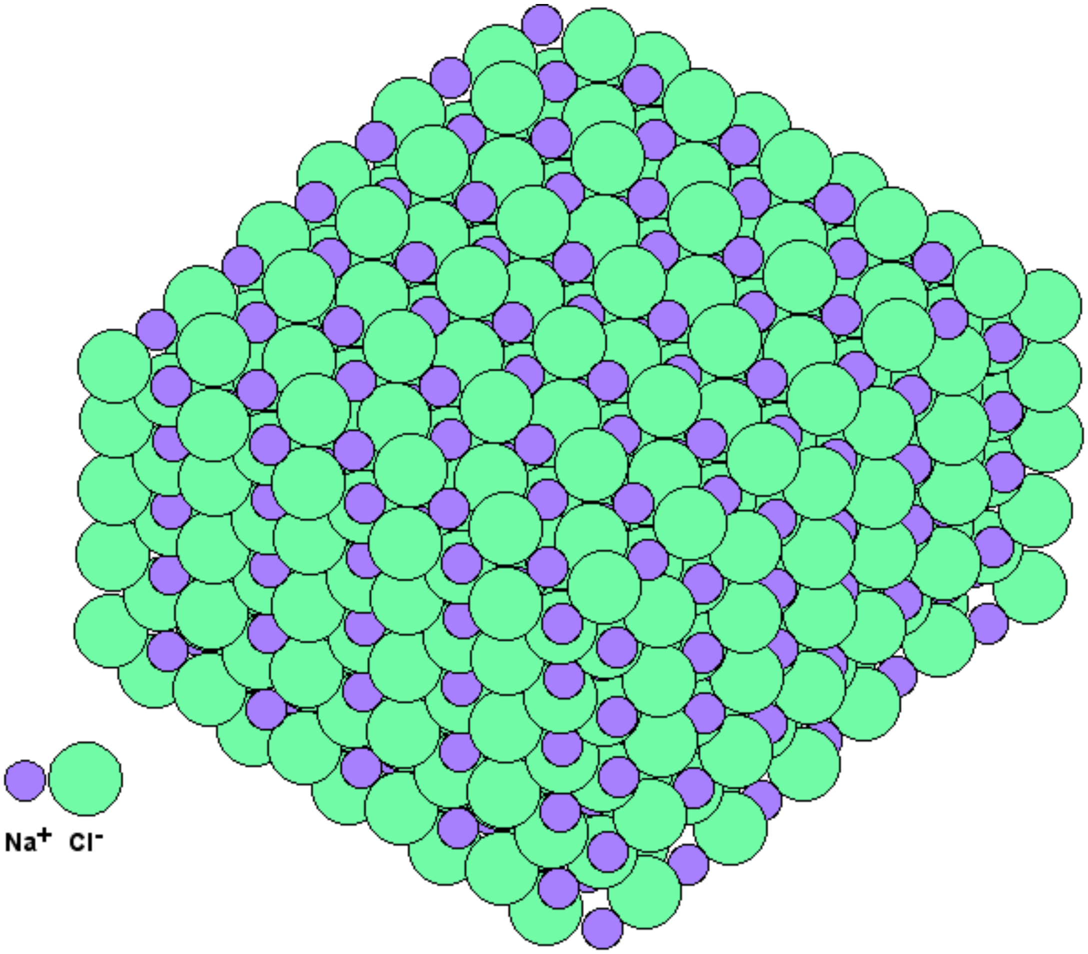

Figure 9.5. Sodium chloride crystal.

When an ionic compound dissolves in water (Figure 9.6), the compound dissociates into anions and cations. The polar ends of water molecules are attracted to the anions and cations in the crystal. Water molecules pull anions and cations out of their proper places in the crystal, eventually surrounding and isolating the ions. Once one ion has been removed, additional locations become available for more water molecules; the process continues until all of the ions are surrounded and isolated by water molecules. The ionic compound is said to ***dissociate*** into individual cations and anions.

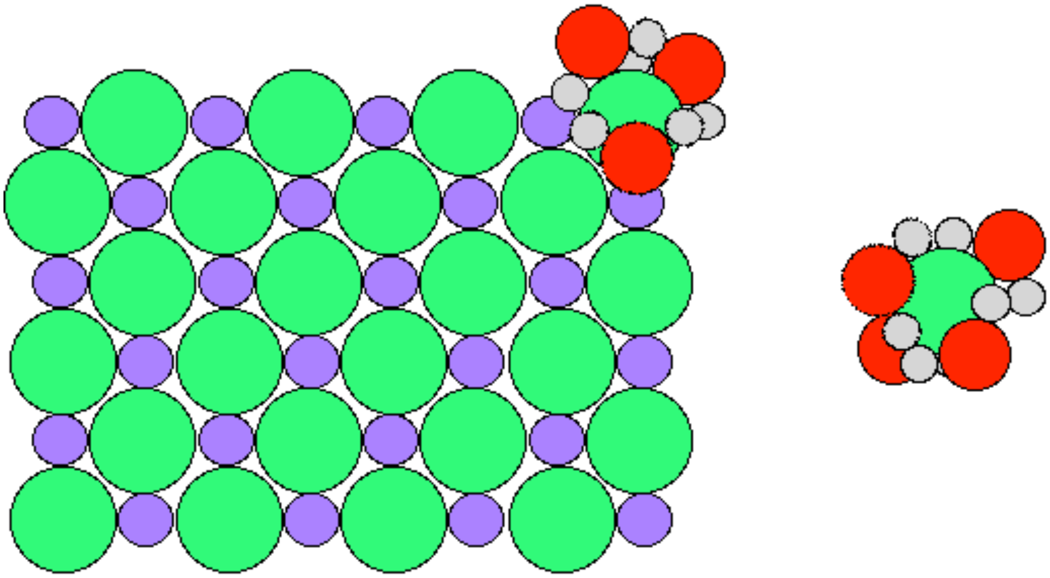

Figure 9.6. Water molecules are surrounding and removing a chloride ion (green) from a sodium chloride crystal. Sodium ions are purple, hydrogen atoms are grey, and oxygen atoms are red.

If the ionic compound dissociates almost 100% in water, then it is called a ***strong electrolyte***. Other ionic compounds dissociate much less than 100% in water, and are called ***weak electrolytes***. Sodium chloride is an example of a strong electrolyte, while silver chloride is an example of a weak electrolyte.

Electrolytes are one of two classes of substances that dissolve in water. The other class is ***nonelectrolytes***, compounds that dissolve without dissociating into cations and anions. Nonelectrolytes include substances like alcohols (CH_3CH_2OH), and sugars ($C_6H_{12}O_6$), among others. Students need to distinguish between electrolytes and nonelectrolytes. Nonelectrolytes have individual, intact molecules dispersed throughout the water solution and tend to be nonionic compounds (compounds formed exclusively from non-metallic elements). Electrolytes don't have individual, intact molecules; instead they have anions and cations dispersed throughout the water solution. All ionic compounds are also electrolytes.

We can distinguish between electrolytes and nonelectrolytes by measuring the ***conductivity*** of water; how easily an electric current passes through water. Pure water has very low conductivity; practically no current flows through pure water. A water solution of a nonelectrolyte, such as sugar, also has low conductivity, while water solutions of a strong electrolyte, such as salt, has a high conductivity. Weak electrolytes exhibit intermediate conductivity.

One convenient and simple conductivity cell can be made from a beaker, a battery, a light bulb, two strips of aluminum foil, and three wires (Figure 9.7). The test solution is put into the beaker. If the solution is a nonelectrolyte, the light bulb won't glow. If the test solution is a weak electrolyte, the bulb will glow relatively faintly, while if it is a strong electrolyte, the bulb will glow brightly.

NOTE: Mixing water and electricity is an extremely dangerous activity that can result in serious injury or death. Anyone making or using such a device does so at his or her own peril. I am not responsible for the consequences of your actions.

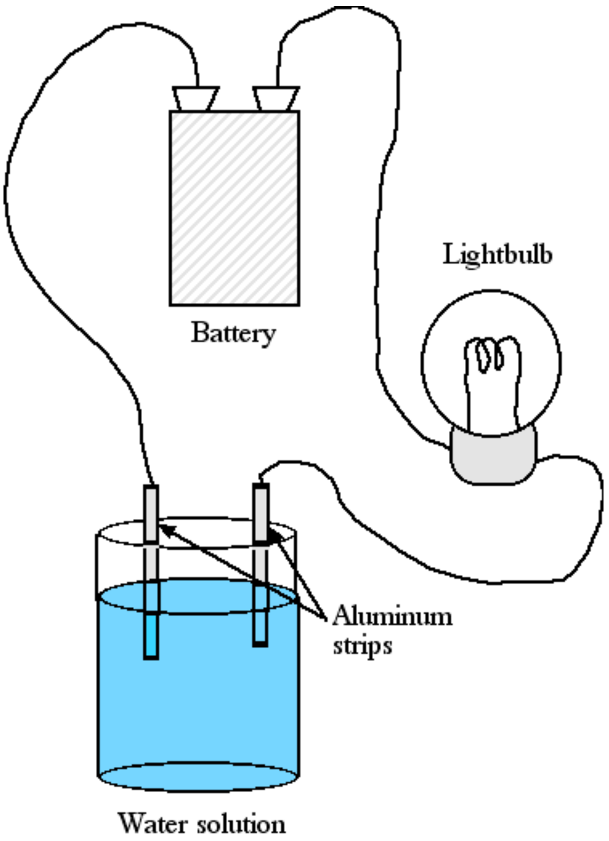

Figure 9.7. Simple conductivity cell.

Using results from conductivity and other types of testing, chemists have developed the solubility rules described above. The stronger the electrolyte, the more soluble the compound will be in water, while weaker electrolytes have lower solubilities.

Nonionic compounds that are polar molecules dissolve in water (a polar solvent). Mutual attractions between the polar ends of the solute molecules with the polar ends of water molecules are responsible for the solubility of polar nonionic compounds. Hydrogen bonding, where appropriate, contributes to the solubility of many nonionic compounds in water, with sugars and alcohols two examples of hydrogen bonding.

Chemists frequently use the expression “Like dissolves like” to explain the solubility (or lack of solubility) of solutes in solvents. Ionic or polar solutes are soluble in polar solvents, but are insoluble in nonpolar solvents. Nonpolar solutes are soluble in nonpolar solvents, but are insoluble in polar solvents. Of course, on an absolute scale, everything is soluble in everything else, even if it is only one molecule of solute in 100,000,000 liters of solvent.

Increasing dissolution rate.

Generally, there are three methods for increasing the dissolution rate. First, you can reduce the particle size by grinding the solid solute into smaller particles. As the particle size shrinks, the total surface area exposed to the solvent increases, and therefore the particles dissolve faster. Second, you can stir or agitate the solution. Stirring uniformly disperses the dissolved solute throughout the solvent, and brings fresh portions of solvent into contact with undissolved solute. Third, you can heat the solution. Generally, solutes are more soluble at high temperature than at low temperature, although there are certain exceptions (gases in water, for example).

Sometimes nothing can be done to increase the dissolution rate, and you simply have to be patient.

Vocabulary. The following terms are defined and explained in the text. Make sure that you are familiar with the meanings of the terms as used in chemistry. Understand that you may have been given incomplete or mistaken meanings for these terms in earlier courses. The meanings given in the text are correct and proper.

Solution
Solute
Solvent
Miscible
Unsaturated
Saturated
Supersaturated
Precipitate
Soluble
Insoluble
Strong electrolyte
Weak electrolyte
Non-electrolyte
Conductivity
Concentration
Molarity
Dissociate
Percent by mass-volume (%m/v)
Percent by mass (%m/m)
Percent by volume (%v/v)

Solubility: Predict whether the following salts are soluble or insoluble in water. Remember to use the solubility rules *in order*, and the first rule that applies determines the solubility of the compound!

1. $CuCl_2$
2. $(NH_4)_2SO_4$
3. Li_3PO_4
4. $Ca(OH)_2$
5. $PbSO_4$
6. MnS
7. $CaCO_3$
8. $Sr(CH_3CO_2)_2$
9. $Fe(OH)_3$
10. CCl_4

Concentration: Perform the indicated calculations.

1. A sodium hydroxide solution is 25% by weight sodium hydroxide dissolved in water. What mass of sodium hydroxide is present in 100 mL of this solution, if the density of the solution is 1.25 g/mL?
2. What is the molarity of the solution described in #1, above?
3. 15.00 grams of glucose ($C_6H_{12}O_6$) are dissolved in sufficient water to produced 50.00 mL of solution. What is the molarity of the glucose?
4. A solution is prepared by combining 12.45 grams of sodium chloride, 7.09 grams of sodium sulfate, 5.00 grams of calcium nitrate, and 55.00 grams of water. What is the % m/m of all components of this solution?
5. A solution is labeled "0.1588 M NaCl". How many moles of sodium chloride are in 25.00 mL of this solution? How many grams of sodium chloride are present in 25.00 mL of this solution?
6. 75.00 mL of 0.0500 M sodium nitrate solution is diluted to a final volume of 250.00 mL with deionized water. What is the final concentration of sodium nitrate in the diluted solution?
7. 25.00 mL of 15 M acetic acid is diluted to a final volume of 100.00 mL with deionized water. How many grams of acetic acid are in the diluted solution?
8. 25.00 mL of ethyl alcohol is diluted in sufficient water to make a solution having a volume of 65.00 mL. What is the volume percentage of ethyl alcohol (% v/v) in this solution?
9. Which solution contains the greater mass of glucose: 100.00 mL of 0.1000 M glucose or 100.00 mL of 10.00% m/m glucose solution with a density of 1.156 g/mL?
10. What is the molarity of glucose in the 10.00% m/m glucose solution described in #9 (above)?

Solution stoichiometry. Perform the indicated calculations.

1. A solution is prepared by dissolving 10.8 g of ammonium sulfate in enough water to produce 100.0 mL of stock solution. A 10.00 mL portion of the stock solution is added to 50.00 mL of water. Calculate the molar concentration of ammonium ions and sulfate ions.

2. The concentration of magnesium ions and calcium ions in a water sample are 30 and 75 mg/L respectively. What are their respective molar concentrations?

3. If fluoride ion concentration exceeds 2 – 3 mg/L, a brown mottling of teeth can occur. What are the corresponding molarities?

4. What mass of AgBr is produced when 100.0 mL of 0.1250 M $AgNO_3$ is added to 20.0 mL of 1.00 M NaBr?

5. What volume of 0.100 M NaOH is required to precipitate all of the nickel(II) ion from 150.0 mL of 0.250 M $Ni(NO_3)_2$?

6. What mass of barium sulfate is produced when 50.0 mL of 0.100 M barium chloride is mixed with 50.0 mL of 0.050 M iron(III) sulfate?

7. What volume of the following acids are needed to react completely with 50.00 mL of 0.200 M NaOH?

   a. 0.100 M HCl
   b. 0.100 M H_2SO_3 (2 acid hydrogens)
   c. 0.150 M H_3PO_4 (3 acid hydrogens)
   d. 0.075 M HNO_3
   e. 0.250 M $HC_2H_5O_2$ (1 acid hydrogen)
   f. 0.0375 M H_2SO_4 (2 acid hydrogens)

8. Sodium hydroxide solution is standardized by titrating a pure sample of potassium hydrogen phthalate (KHP), an acid having one acidic hydrogen and a molecular weight of 204.22. If it takes 20.46 mL of sodium hydroxide solution to titrate a 0.1037 g sample of KHP, what is the molar concentration of sodium hydroxide?

9. A 10.00 mL sample of battery acid (sulfuric acid) requires 55.89 mL of 1.89 M sodium hydroxide for complete neutralization. What is the molarity of the battery acid?

10. A 50.00 mL portion of ammonia solution is titrated with HCl. The reaction is:

$$NH_3 + H^+ \rightarrow NH_4^+$$

If the titration consumes 39.11 mL of 0.0889 M HCl, what is the original concentration of ammonia?

Answers:

Solubility:

1. $CuCl_2$

   The ions produced are Cu^{+2} and Cl^{-}. Copper ion does not have a specific solubility rule, nor does it appear as an important exception. Chloride salts are soluble according to rule 3. We predict that this salt is soluble.

2. $(NH_4)_2SO_4$

   The ions produced are ammonium (NH_4^{+}) and sulfate (SO_4^{-2}). All ammonium salts are soluble (by rule 1). We predict that this salt is soluble.

3. Li_3PO_4

   The ions produced are lithium (Li^{+}) and phosphate (PO_4^{-3}). All salts containing a group 1A metal are soluble, and lithium is a group 1A metal, so by rule 1 this salt is soluble.

4. $Ca(OH)_2$

   Calcium ion (Ca^{+2}) and hydroxide ion (OH^{-}) are produced from this compound. There is no specific rule concerning calcium ion, although it does form some important exceptions. Hydroxide salts are generally insoluble (rule 5), but calcium ion is sometimes an important exception in this solubility rule. We can't make a firm prediction based on the solubility rules, but calcium hydroxide may be soluble in water.

5. $PbSO_4$

   Lead ion (Pb^{+2}) and sulfate (SO_4^{-2}) are present in this salt. There is no specific rule concerning the solubility of lead compounds, although lead does form some important exceptions. Sulfate salts are generally soluble (rule 4), however lead is an important exception to this rule. We predict that this salt is insoluble.

6. MnS

   Manganese ion (Mn^{+2}) and sulfide ion (S^{-2}) are present in this salt. There is no specific rule concerning manganese ion. Sulfide salts are generally insoluble (rule 6). While there are exceptions to this rule, manganese is not listed as one of these exceptions. We predict that this salt is insoluble.

7. $CaCO_3$

   The ions involved are calcium ion (Ca^{+2}) and carbonate ion (CO_3^{-2}). While calcium ion is an important exception to certain solubility rules, there is no specific rule concerning calcium ion. Carbonate salts are generally insoluble (rule 7), with the exception of ammonium and group 1A metal cations. Since calcium is neither ammonium nor a group 1A metal cation, we predict that this salt is insoluble.

8. $Sr(CH_3CO_2)_2$

   Strontium ion (Sr^{+2}) and acetate ion ($CH_3CO_2^-$) are present in this compound. There is no specific rule concerning strontium ion, although it does appear in certain exceptions. By rule 2, we predict that this salt is soluble.

9. $Fe(OH)_3$

   Iron(III) ion (Fe^{+3}) and hydroxide ion (OH^-) are present in this salt. There is no specific rule concerning iron, and it does not form an important exception to any solubility rule. Hydroxide salts are insoluble (rule 5). Therefore, this salt is insoluble.

10. CCl_4

    If you started to solve this problem by writing the ions as C^{+4} and Cl^-, then shame on you! This is not an ionic compound and therefore cannot be a salt. This compound is formed of the nonmetals carbon and chloride. Review Chapter 4 concerning the identification of ionic compounds and nonionic compounds.

Concentration.

1. 100 mL x 1.25 g/mL = 125 grams solution.

   125 grams solution x (25/100) = 31.25 grams sodium hydroxide.

2. F.W. sodium hydroxide (NaOH) = 40.00 g/mole.

   31.25 grams sodium hydroxide/40.00 g/mole = 0.78125 moles sodium hydroxide.

   0.78125 moles sodium hydroxide / 0.100 L = 7.8125 M (7.8 M)

3. MW glucose = 180.1788 g/mole.

   15.00 g glucose / 180.1788 g/mole = 0.08325 moles glucose.

   0.08325 moles glucose / 0.05000 L = 1.665 M

4. Total mass = 12.45 g + 7.09 g + 5.00 g + 55.00 g = 79.54 g.

   % sodium chloride = (12.45 g / 79.54 g) x 100 = 15.65%

   % sodium sulfate = (7.09 g / 79.54 g) x 100 = 8.91%

   % calcium nitrate = (5.00 g / 79.54 g) x 100 = 6.29%

   % water = (55.00 g / 79.54 g) x 100 = 69.15%

5. 0.1588 M = 0.1588 moles/liter = 0.1588 moles/1000 mL solution.

   (0.1588 moles/1000 mL) x 25.00 mL = 0.003970 moles NaCl

   0.003970 moles sodium chloride x 58.44 g/mole = 0.2320 g NaCl

6. $M_iV_i = M_fV_f$

   75.00 mL x 0.0500 M = M_f x 250.00 mL

   M_f = 0.0150 M

7. 0.02500 L x 15 moles/liter = 0.3750 moles acetic acid

   0.3750 moles x 60.05 g/mole = 22.52 grams acetic acid.

8. (25.00 mL ethyl alcohol / 65.00 mL total solution) x 100 = 38.46 % v/v

9. 0.1000 L x 0.1000 M = 0.01000 moles glucose.

   0.01000 moles glucose x 180.1788 g/ mole = 1.802 grams glucose

   100.00 mL x 1.156 g/mL = 115.6 g total solution.

   115.6 g total solution x (10.00/100.0) = 1.156 grams glucose.

   The 0.1000 M glucose contains more glucose than does the
   10.00% m/m glucose solution.

10. (1.156 g glucose / 180.1788 g/mole) = 0.006416 moles glucose.

0.006416 moles / 0.10000 L = 0.06416 M.

Solution stoichiometry.

1. FW $(NH_4)_2SO_4$ = 132.14

$$10.8g \text{ ammonium sulfate } \times \frac{1 \text{ mole}}{132.14g} = 0.0817moles$$

$$\frac{0.0817moles}{0.1000L} = 0.817M$$

$$10.00mL \times \frac{0.817M}{60.00mL} = 0.136M \ (NH_4)_2SO_4$$

$$SO_4^{-2} \text{ molarity} = 0.136\text{M}$$

$$NH_4^{+} \text{ molarity} = 2 \text{ x } 0.136\text{M} = 0.272\text{M}$$

2. Mg^{+2}: 30 mg/L = 0.030 g/L
   0.030 g Mg^{+2}/ 24.305g/mol = 0.0012 moles
   0.0012 moles/ 1 liter = 0.0012M

   Ca^{+2}: 75 mg/L = 0.075 g/L
   0.075 g Ca^{+2}/40.078g/mole =0.0019 moles
   0.0019 moles/ 1 liter = 0.0019M

3. 2 mg F = 0.002 g F = 0.000 1 mole F
   3 mg F = 0.003 g F = 0.000 16 mole F

   2 – 3 mg/L = 0.000 1 M – 0.000 16 M

4. $AgNO_3 + NaBr \rightarrow AgBr + NaNO_3$

   0.1000 L x 0.1250 M = 0.0125 moles $AgNO_3$

   0.0200 L x 1.00 M = 0.0200 moles NaBr

   $AgNO_3$ limits reaction; only 0.0125 moles of AgBr can be produced.

   F.W. AgBr = 187.77

   0.0125 moles x 187.77 g/mole = 2.35 g AgBr

5. $Ni^{+2} + 2OH^{-} \rightarrow Ni(OH)_2$

   $0.1500\ L \times 0.250\ M = 0.0375$ moles $Ni(NO_3)_2 = 0.0375$ moles Ni^{+2}

$$0.0375 \text{ moles } Ni^{+2} \text{ x } \frac{2\ OH^{-}}{1\ Ni^{+2}} = 0.075 \text{ moles } OH^{-}$$

$$\frac{0.075 \text{ moles } OH^{-}}{0.100M \text{ NaOH}} = 0.75 \text{ L NaOH} = 750 \text{ mL}$$

6. $3BaCl_2 + Fe_2(SO_4)_3 \rightarrow 3BaSO_4 + 2FeCl_3$

   0.0500 L x 0.100 M = 0.005 00 moles barium chloride

   0.0500 L x 0.050 M = 0.0025 moles iron(III) sulfate

$$0.005\ 00 \text{ moles barium chloride } \times \frac{1 \text{ iron(III) sulfate}}{3 \text{ barium chloride}} = 0.00167 \text{ moles iron(III) sulfate}$$

   barium chloride limiting – when it is all consumed, iron(III) sulfate remains.

   0.00500 moles barium chloride produces 0.00500 moles barium sulfate

   F.W. barium sulfate = 233.39

   0.00500 moles x 233.39 g/mole = 1.17 grams barium sulfate

7. 0.05000 L x 0.200 M = 0.0100 moles NaOH.

$$0.0100 \text{ moles NaOH } \times \frac{1 \text{ HCl}}{1 \text{ NaOH}} = 0.0100 \text{ moles HCl}$$

   a.

$$\frac{0.0100 \text{ moles HCl}}{0.100 \text{ M HCl}} = 0.100 \text{ L (100 mL)}$$

b.

$$0.0100 \text{ moles NaOH } \times \frac{1 \text{ H}_2\text{SO}_3}{2 \text{ NaOH}} = 0.00500 \text{ moles H}_2\text{SO}_3$$

$$\frac{0.00500 \text{ moles H}_2\text{SO}_3}{0.100 \text{ M H}_2\text{SO}_3} = 0.0500 \text{ L (50.0 mL)}$$

c.

$$0.0100 \text{ moles NaOH } \times \frac{1 \text{ H}_3\text{PO}_4}{3 \text{ NaOH}} = \ 0.003\ 33 \text{ moles H}_3\text{PO}_4$$

$$\frac{0.003\ 33 \text{ moles H}_3\text{PO}_4}{0.150 \text{ M H}_3\text{PO}_4} = 0.0222 \text{ L (22.2 mL)}$$

d.

$$0.0100 \text{ moles NaOH } \times \frac{1 \text{ HNO}_3}{1 \text{ NaOH}} = 0.0100 \text{ moles HNO}_3$$

$$\frac{0.0100 \text{ moles HNO}_3}{0.075 \text{ M HNO}_3} = 0.13 \text{ L (130 mL)}$$

e.

$$0.0100 \text{ moles NaOH } \times \frac{1 \text{ HC}_2\text{H}_3\text{O}_2}{1 \text{ NaOH}} = 0.0100 \text{ moles HC}_2\text{H}_3\text{O}_2$$

$$\frac{0.0100 \text{ moles HC}_2\text{H}_3\text{O}_2}{0.250 \text{ M HC}_2\text{H}_3\text{O}_2} = 0.040 \text{ L (40.0 mL)}$$

f.

$$0.0100 \text{ moles NaOH } \times \frac{1 \text{ H}_2\text{SO}_4}{2 \text{ NaOH}} = 0.00500 \text{ moles H}_2\text{SO4}$$

$$\frac{0.00500 \text{ moles H}_2\text{SO}_4}{0.0375 \text{ M H}_2\text{SO}_4} = 0.133 \text{ L (133 mL)}$$

8. 0.1037 g KHP / 204.22 g/mole = 0.000 507 8 moles KHP = 0.000 507 8 moles NaOH

0.000 507 8 moles NaOH/0.02046 L = 0.02482 M

9. $H_2SO_4 + 2\ NaOH \rightarrow 2H_2O + Na_2SO_4$

0.05589 L x 1.89 M = 0.106 moles NaOH

$$0.106 \text{ moles NaOH} \times \frac{1\ H_2SO_4}{2\ NaOH} = 0.053 \text{ moles } H_2SO_4$$

$$\frac{0.053 \text{ moles } H_2SO_4}{0.01000 \text{ L}} = 0.53 \text{ M}$$

10. 0.03911 L x 0.0889 M = 0.003 48 moles HCl = 0.003 48 moles NH3

$$\frac{0.003\ 48 \text{ moles } NH_3}{0.05000 \text{ L}} = 0.0695 \text{ M}$$