Solutions to the Extra Problems for Chapter 11

1. a. The CH_5P became CH_6P^+, which means it gained an H^+, so $\underline{CH_5P}$ is the base. The $C_4H_6O_2$ lost an H^+ to become $C_4H_5O_2^-$, so $\underline{C_4H_6O_2}$ is the acid.

b. The $Si_2H_2O_5$ lost an H^+ to become $Si_2HO_5^-$, so $\underline{Si_2H_2O_5}$ is the acid. The $C_8H_{18}BaO_2$ gained an H^+ to become $C_8H_{19}BaO_2^+$, so $\underline{C_8H_{18}BaO_2}$ is the base.

2. The reaction is $\underline{HI + H_2O \rightarrow H_3O^+ + I^-}$. If HI is going to force H_2O to act as a base, it will donate an H^+ to the H_2O. That will leave I^-, and adding an H^+ to water makes H_3O^+.

3. a. The reaction is $\underline{HCl + NH_3 \rightarrow Cl^- + NH_4^+}$. You are supposed to recognize that NH_3 is a base. You are also supposed to recognize HCl as an acid. That means it will donate an H^+, leaving only Cl^-. When NH_3 accepts an H^+, the result will be NH_4^+.

b. The reaction is $\underline{3H_2CrO_4 + 2Al(OH)_3 \rightarrow 6H_2O + Al_2(CrO_4)_3}$. This is an acid and an ionic base, so we can use acid + base makes water + salt. When the H_2CrO_4 donates its H^+ ions, CrO_4^{2-} is left behind. When the H^+ ions combine with the hydroxides from the base, Al^{3+} is left behind. That means the salt comes from Al^{3+} and CrO_4^{2-}, so it's $Al_2(CrO_4)_3$. The unbalanced equation is:

$$H_2CrO_4 + Al(OH)_3 \rightarrow H_2O + Al_2(CrO_4)_3$$

To balance, we start by skipping the H's. To balance the Cr's we need to multiply H_2CrO_4 by 3. To balance the Al's we need to multiply $Al(OH)_3$ by 2:

$$3H_2CrO_4 + 2Al(OH)_3 \rightarrow H_2O + Al_2(CrO_4)_3$$

Now we have to look at the H's or O's. Since H appears only once on the products side, that's what we should look at first. There are 3×2 + 2×3 = 12 H's on the reactants side, so we have to multiply H_2O by 6:

$$3H_2CrO_4 + 2Al(OH)_3 \rightarrow 6H_2O + Al_2(CrO_4)_3$$

Now everything is balanced.

c. The reaction is $\underline{H_3PO_4 + 3NaOH \rightarrow 3H_2O + Na_3PO_4}$. The acid can donate three H^+ ions, leaving PO_4^{3-} behind. When the ions combine with the hydroxide from the base, Na^+ is left over. The salt, therefore, is formed by Na^+ and PO_4^{3-}, so it is Na_3PO_4.

$$H_3PO_4 + NaOH \rightarrow H_2O + Na_3PO_4$$

To balance, we skip the H's. The P's are balanced, and we will skip the O's. To balance the Na's, we need to multiply NaOH by 3:

$$H_3PO_4 + 3NaOH \rightarrow H_2O + Na_3PO_4$$

Now we have to deal with the O's or H's. There are a total of 1×3 + 3×1 = 6 H's on the reactants side, so we need to multiply H_2O by 3:

$$H_3PO_4 \ + \ 3NaOH \ \rightarrow \ 3H_2O \ + \ Na_3PO_4$$

Now everything is balanced.

4. <u>It will require 0.0067 L (6.7 mL)</u>. This is just a stoichiometry problem where we need to figure out how much acid is required to neutralize a specific amount of base. To figure that out, we first must determine the chemical equation. HCl is the acid, and when it donates its H^+, the only thing left will be Cl^-. When the H^+ ions add to the OH^- ions in $Ca(OH)_2$ to make water, the Ca^{2+} ion will be left over. Thus, the salt that is produced in the reaction will be made from Ca^{2+} and Cl^-. Switching the numbers and dropping the charges gives us $CaCl_2$. That means the equation is:

$$HCl \ + \ Ca(OH)_2 \ \rightarrow \ H_2O \ + \ CaCl_2$$

To balance, we skip the H's. To balance the Cl's, we have to multiply HCl by 2. The Ca's are balanced, so we go back to the H's. To balance them, we have to multiply H_2O by 2:

$$2HCl \ + \ Ca(OH)_2 \ \rightarrow \ 2H_2O \ + \ CaCl_2$$

Now we can figure out how much HCl we need. First, let's use the concentration and volume to determine the moles of $Ca(OH)_2$. To do that, however, the volume must be converted to 0.2500 liters.

$$0.010\frac{\text{moles}}{\cancel{\text{L}}} \times 0.2500\,\cancel{\text{L}} = 0.0025 \text{ moles}$$

We can now convert that to moles of HCl:

$$\frac{0.0025\ \cancel{\text{moles Ca(OH)}_2}}{1} \times \frac{2 \text{ moles HCl}}{1\ \cancel{\text{mole Ca(OH)}_2}} = 0.0050 \text{ moles HCl}$$

Now we can use the definition of molarity to determine the volume of HCl:

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$0.75\,\frac{\text{moles}}{\text{L}} = \frac{0.0050 \text{ moles}}{\text{liters of solution}}$$

Multiplying both sides of the equation by "liters of solution" and dividing both sides by 0.75 moles/liter gives us:

$$\text{liters of solution} = \frac{0.0050\ \cancel{\text{moles}}}{.75\,\frac{\cancel{\text{moles}}}{\text{L}}} = 0.0067 \text{ L}$$

5. It will require 160 g. This is another stoichiometry problem where we need to figure out how much base is required to neutralize a specific amount of acid. To figure that out, we first must determine the chemical equation. H_2SO_4 is the acid, and when it donates its two H^+ ions, the only thing left will be SO_4^{2-}. When the H^+ ions add to the OH^- ions in $Al(OH)_3$ to make water, the Al^{3+} ion will be left over. Thus, the salt that is produced in the reaction will be made from Al^{3+} and SO_4^{2-}. Switching the numbers and dropping the charges gives us $Al_2(SO_4)_3$. That means the reaction is:

$$H_2SO_4 + Al(OH)_3 \rightarrow H_2O + Al_2(SO_4)_3$$

To balance, we skip the H's. To balance the S's, we have to multiply H_2SO_4 by 3. To balance the Al's, we have to multiply $Al(OH)_3$ by 2.

$$3H_2SO_4 + 2Al(OH)_3 \rightarrow H_2O + Al_2(SO_4)_3$$

There are now 3×2 + 2×3 = 12 H's on the reactants side, so we need to multiply H_2O by 6:

$$3H_2SO_4 + 2Al(OH)_3 \rightarrow 6H_2O + Al_2(SO_4)_3$$

Now we have a balanced chemical equation, so we can do stoichiometry. First, let's use the concentration and volume to determine the moles of H_2SO_4:

$$5.6\,\frac{\text{moles}}{\cancel{\text{L}}} \times 0.5500\,\cancel{\text{L}} = 3.1 \text{ moles}$$

We can now convert that to moles of $Al(OH)_3$:

$$\frac{3.1\,\cancel{\text{moles } H_2SO_4}}{1} \times \frac{2 \text{ moles } Al(OH)_3}{3\,\cancel{\text{moles } H_2SO_4}} = 2.1 \text{ moles } Al(OH)_3$$

This question asks for grams, so we just convert that to grams using the molar mass of $Al(OH)_3$.

$$\frac{2.1\,\cancel{\text{moles } Al(OH)_3}}{1} \times \frac{78.01 \text{ g } Al(OH)_3}{1\,\cancel{\text{mole } Al(OH)_3}} = 160 \text{ g } Al(OH)_3$$

6. The concentration is 0.135 M. The endpoint tells us that enough base has been added to neutralize the acid. Once we determine the chemical equation, we can use that fact to get the concentration of the acid. $HClO_3$ donates an H^+ ion, leaving ClO_3^-. When those ions add to the hydroxide ion from the base, Na^+ remains. That means the salt is $NaClO_3$.

$$HClO_3 + NaOH \rightarrow H_2O + NaClO_3$$

The equation is already balanced, so we can do the stoichiometry. We can get the moles of base used by multiplying concentration (2.16 M) by volume in liters (0.00718 L):

$$2.16\ \frac{\text{moles}}{\cancel{\text{L}}} \times 0.00718\ \cancel{\text{L}} = 0.0155\ \text{moles}$$

We can now use the chemical equation to determine how many moles of acid were in the solution:

$$\frac{0.0155\ \cancel{\text{moles NaOH}}}{1} \times \frac{1\ \text{mole HClO}_3}{1\ \cancel{\text{mole NaOH}}} = 0.0155\ \text{moles HClO}_3$$

We were told the volume of the $HClO_3$ solution (115.0 mL). If we convert it to liters (0.1150 L), we can use the definition of molarity to determine the concentration of the acid.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$\text{Molarity} = \frac{0.0155\ \text{moles}}{0.1150\ \text{L}} = 0.135\ \text{M}$$

7. <u>The concentration is 3.19 M</u>. The endpoint tells us that enough acid has been added to neutralize the base. Once we determine the chemical equation, we can use that fact to get the concentration of the base. H_2SO_4 donates two H^+ ions, leaving SO_4^{2-}. When those ions add to the hydroxide ion from the base, K^+ remains. That means the salt is K_2SO_4.

$$H_2SO_4 \ + \ KOH \ \rightarrow \ H_2O \ + \ K_2SO_4$$

To balance the equation, we need to multiply KOH by 2 and H_2O by 2:

$$H_2SO_4 \ + \ 2KOH \ \rightarrow \ 2H_2O \ + \ K_2SO_4$$

Now we can do the stoichiometry. We can get the moles of acid used by multiplying concentration (1.34 M) by volume in liters (0.0178 L):

$$1.34\ \frac{\text{moles}}{\cancel{\text{L}}} \times 0.0178\ \cancel{\text{L}} = 0.0239\ \text{moles}$$

We can now use the chemical equation to determine how many moles of acid were in the solution:

$$\frac{0.0239\ \cancel{\text{moles H}_2\text{SO}_4}}{1} \times \frac{2\ \text{moles KOH}}{1\ \cancel{\text{mole H}_2\text{SO}_4}} = 0.0478\ \text{moles KOH}$$

We were told the volume of the KOH solution (0.0150 L), so we can use the definition of molarity to determine the concentration of the acid.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$\text{Molarity} = \frac{0.0478 \text{ moles}}{0.0150 \text{ L}} = 3.19 \text{ M}$$

8. <u>You should take 0.063 L of the concentrated solution and add it to enough water to make 750.0 mL</u>. We first need to figure out how many moles of acetic acid we need in order to make the solution. Since we know the volume (0.7500 L) and the concentration (1.5 M) we want, we can multiply the two to get the moles we need:

$$1.5 \frac{\text{moles}}{\cancel{\text{L}}} \times 0.7500 \cancel{\text{L}} = 1.1 \text{ moles}$$

Where do those moles come from? They come from the concentrated solution. We know the molarity of the concentrated solution, so we can use that to determine the volume that we need to give us that many moles:

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$17.4 \text{ M} = \frac{1.1 \text{ moles}}{\text{liters of solution}}$$

Multiplying both sides of the equation by "liters of solution" and then dividing by 17.4 M gives us:

$$\text{liters of solution} = \frac{1.1 \cancel{\text{moles}}}{17.4 \frac{\cancel{\text{moles}}}{\text{L}}} = 0.063 \text{ L}$$

9. The most the student could make is <u>0.17 L of solution</u>. If the student uses all of the original solution, we know how many moles he has, because we know the volume of the original solution (0.0500 L) and the concentration (5.0 M):

$$5.0 \frac{\text{moles}}{\cancel{\text{L}}} \times 0.0500 \cancel{\text{L}} = 0.25 \text{ moles}$$

Now that we know the number of moles, we can use the definition of molarity to figure out the volume of solution that would give that number of moles a concentration of 1.5 M:

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$1.5 \text{ M} = \frac{0.25 \text{ moles}}{\text{liters of solution}}$$

Multiplying both sides of the equation by "liters of solution" and then dividing by 1.5 M gives us:

$$\text{liters of solution} = \frac{0.25\ \cancel{\text{moles}}}{1.5\ \frac{\cancel{\text{moles}}}{\text{L}}} = 0.17\ \text{L}$$

10. It was 8.6 M. We don't know the initial number of moles, because we don't have a concentration for the original solution. However, we do know the final number of moles, because we know the volume of the final solution (0.250 L) and the concentration of the final solution (1.7 M). So, the total number of moles is:

$$1.7\ \frac{\text{moles}}{\cancel{\text{L}}} \times 0.250\ \cancel{\text{L}} = 0.43\ \text{moles}$$

Those moles had to come from the original solution. Since we know the volume of the original solution, we can now solve for the concentration of the original solution

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$1.5\ \text{M} = \frac{0.43\ \text{moles}}{0.0500\ \text{L}} = 8.6\ \text{M}$$