General Chemistry I
PLI Sheet 5 AK
June 9, 2021

1. **Quinone, which is used in the dye industry and in photography, is an organic compound containing only C, H, and O. What is the empirical formula of the compound if you find that 0.105 g of the compound gives 0.257 g of $CO_2$ and 0.0350 g of $H_2O$ when burned completely? Given a molecular weight of approximately 108 g/mol, what is its molecular formula?**

Another combustion analysis problem!

$$0.0350\ g\ H_2O * \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} * \frac{2\ mol\ H}{1\ mol\ H_2O} = 0.00388\ mol\ H * \frac{1.008\ g\ H}{1\ mol\ H} = 0.00392\ g\ H$$

$$0.257\ g\ CO_2 * \frac{1\ mol\ CO_2}{44.01\ g\ CO_2} * \frac{1\ mol\ C}{1\ mol\ CO_2} = 0.005840\ mol\ C * \frac{12.011\ g\ C}{1\ mol\ C} = 0.0701\ g\ C$$

$$Mass\ O = Mass\ sample - Mass\ C - Mass\ H = 0.105\ g - 0.0701\ g - 0.00392\ g = 0.03098\ g\ O$$

$$0.03098\ g\ O * \frac{1\ mol\ O}{15.9994\ g\ O} = 0.001936\ mol\ O$$

Take mole ratios to find empirical formula!

$$C: \frac{0.005840}{0.001936} = 3$$

$$H: \frac{0.00388}{0.001936} = 2$$

$$O: \frac{0.001936}{0.001936} = 1$$

Empirical formula: $C_3H_2O$

$$\frac{Molar\ mass}{Empirical\ formula\ mass} = \frac{108\frac{g}{mol}}{54\frac{g}{mol}} = 2$$

Molecular formula: $C_6H_4O_2$

Balanced equation $C_6H_4O_2 + 6\ O_2 \rightarrow 2\ H_2O + 6\ CO_2$

2. **Aqueous sodium sulfide is mixed with iron(III) chloride to produce iron(III) sulfide and sodium chloride. Please make sure to include all states of matter in the equations.**

   **Balanced chemical equation:**

$$3Na_2S(aq) + 2FeCl_3(aq) \rightarrow Fe_2S_3(s) + 6NaCl(aq)$$

   **Total ionic equation:**

$$6Na^+(aq) + 3S^{2-}(aq) + 2Fe^{3+}(aq) + 6Cl^-(aq) \rightarrow Fe_2S_3(s) + 6Na^+(aq) + 6Cl^-(aq)$$

   **Net ionic equation:**

$$3S^{2-}(aq) + 2Fe^{3+}(aq) \rightarrow Fe_2S_3(s)$$

3. **Balance the following equations:**

   a. $\boldsymbol{NH_4NO_2} \rightarrow \boldsymbol{N_2} + 2\boldsymbol{H_2O}$
   b. $\boldsymbol{P_4O_{10}} + 6\boldsymbol{H_2O} \rightarrow 4\boldsymbol{H_3PO_4}$
   c. $2\boldsymbol{HCl} + \boldsymbol{CaCO_3} \rightarrow \boldsymbol{CaCl_2} + \boldsymbol{CO_2} + \boldsymbol{H_2O}$
   d. $2\boldsymbol{C_6H_{14}} + 19\boldsymbol{O_2} \rightarrow 12\boldsymbol{CO_2} + 14\boldsymbol{H_2O}$

4. **Write the overall balanced, total ionic, and net ionic equations for the following reactions:**

   a. $\boldsymbol{HCH_3CO_2(aq)} + \boldsymbol{NH_3(aq)} \rightarrow$

$$HCH_3CO_2(aq) + NH_3(aq) \rightarrow CH_3CO_2^-(aq) + NH_4^+(aq)$$

   There are no spectator ions ($HCH_3CO_2$ is a weak acid), so the equation above is the overall balanced, total ionic, and net ionic equation.

   b. **Aqueous ammonia reacts with aqueous hydrochloric acid**
   Overall balanced:

$$NH_3(aq) + HCl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$$

   Total ionic:

$$NH_3(aq) + H^+(aq) + Cl^-(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$$

   Net ionic:

$$NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$$

c. **Aqueous solutions of aluminum nitrate and potassium phosphate**

Overall balanced:

$$Al(NO_3)_3(aq) + K_3PO_4(aq) \rightarrow AlPO_4(s) + 3KNO_3$$

Total ionic:

$$Al^{3+}(aq) + 3NO_3^-(aq) + 3K^+(aq) + PO_4^{3-}(aq) \rightarrow AlPO_4(s) + 3K^+(aq) + 3NO_3^-(aq)$$

Net ionic:

$$Al^{3+}(aq) + PO_4^{3-}(aq) \rightarrow AlPO_4(s)$$

5. **The reaction of ethane gas ($C_2H_6$) with chlorine gas produces $C_2H_5Cl$ as its main product (along with HCl). In addition, the reaction invariably produces a variety of other minor products, including $C_2H_4Cl_2$, $C_2H_3Cl_3$, and others. Naturally, the production of these minor products reduces the yield of the main product. Calculate the percent yield of $C_2H_5Cl$ if the reaction of 300 g of ethane with 650 g of chlorine produced 490 g of $C_2H_5Cl$.**

$$C_2H_6 + Cl_2 \rightarrow C_2H_5Cl + HCl$$

The equation is already balanced! Be sure to check though
Next, find the limiting reactant

$$300\ g\ C_2H_6 * \frac{1\ mol\ C_2H_6}{30.07\ g\ C_2H_6} * \frac{1\ mol\ C_2H_5Cl}{1\ mol\ C_2H_6} = 9.977\ mol\ C_2H_5Cl$$

$$650\ g\ Cl_2 * \frac{1\ mol\ Cl_2}{70.906\ g\ Cl_2} * \frac{1\ mol\ C_2H_5Cl}{1\ mol\ Cl_2} = 9.167\ mol\ C_2H_5Cl$$

Since $Cl_2$ leads to a smaller amount of $C_2H_5Cl$, $Cl_2$ is the limiting reactant. To calculate percent yield, determine how much product should have theoretically formed.

$$9.167\ mol\ C_2H_5Cl * \frac{64.51\ g\ C_2H_5Cl}{1\ mol\ C_2H_5Cl} = 591.367\ g\ C_2H_5Cl$$

$$\%\ yield = \frac{Actual\ mass}{Theoretical\ mass} * 100\% = \frac{490\ g}{591.367\ g} * 100\% = 82.9\%$$

6. **A solution is prepared by dissolving 10.8 g ammonium sulfate in enough water to make 100.0 mL of stock solution. A 10.00 mL sample of this stock solution is added to 50.00 mL of water. Calculate the concentration of the overall solution, ammonium ions, and sulfate ions.**

$$10.8\ g\ (NH_4)_2SO_4 * \frac{1\ mol\ (NH_4)_2SO_4}{132.14\ g\ (NH_4)_2SO_4} = 0.0817\ mol\ (NH_4)_2SO_4$$

Find the molarity of the stock solution:

$$Stock\ solution\ molarity = \frac{0.0817\ mol\ (NH_4)_2SO_4}{0.100\ L} = 0.817\ M$$

Next, determine how many moles are extracted in the 10.00 mL sample of the stock solution.

$$0.817\ M = \frac{0.817\ mol\ (NH_4)_2SO_4}{1\ L} * 0.01\ L = 0.00817\ mol\ (NH_4)_2SO_4$$

Since 10.00 mL is added to 50.00 mL of water, the total volume of the new solution is 60.00 mL.

$$[(NH_4)_2SO_4] = \frac{0.00817\ mol\ (NH_4)_2SO_4}{0.06\ L} = 0.136\ M$$

$$[NH_4^+] = \frac{0.00817\ mol\ (NH_4)_2SO_4}{0.06\ L} * \frac{2\ mol\ NH_4^+}{1\ mol\ (NH_4)_2SO_4} = 0.272\ M$$

$$[SO_4^{2-}] = \frac{0.00817\ mol\ (NH_4)_2SO_4}{0.06\ L} * \frac{1\ mol\ SO_4^{2-}}{1\ mol\ (NH_4)_2SO_4} = 0.136\ M$$

7. **A 100.0 mL sample of 0.200 M aqueous potassium hydroxide is mixed with 100.0 mL of 0.200 M aqueous magnesium nitrate.**

   a. **Write a balanced chemical equation, total ionic, and net ionic equation for the reaction.**

$$2KOH(aq) + Mg(NO_3)_2(aq) \rightarrow Mg(OH)_2(s) + 2KNO_3$$

$$2K^+(aq) + 2OH^-(aq) + Mg^{2+}(aq) + 2NO_3^-(aq) \\ \rightarrow Mg(OH)_2(s) + 2K^+(aq) + 2NO_3^-(aq)$$

$$2OH^-(aq) + Mg^{2+}(aq) \rightarrow Mg(OH)_2(s)$$

   b. **What mass of precipitate forms?**

This is a limiting reactant problem!

$$0.200\ M\ KOH = \frac{0.200\ mol\ KOH}{1\ L} * 0.100\ L = 0.02\ mol\ KOH$$

$$0.02\ mol\ KOH * \frac{1\ mol\ Mg(OH)_2}{2\ mol\ KOH} = 0.01\ mol\ Mg(OH)_2$$

$$0.200\ M\ Mg(NO_3)_2 = \frac{0.200\ mol\ Mg(NO_3)_2}{1\ L} * 0.100\ L = 0.02\ mol\ Mg(NO_3)_2$$

$$0.02\ mol\ Mg(NO_3)_2 * \frac{1\ mol\ Mg(OH)_2}{1\ mol\ Mg(NO_3)_2} = 0.02\ mol\ Mg(OH)_2$$

KOH is the limiting reactant, producing 0.01 mol $Mg(OH)_2$

$$0.01\ mol\ Mg(OH)_2 * \frac{58.32\ g\ Mg(OH)_2}{1\ mol\ Mg(OH)_2} = 0.583\ g\ Mg(OH)_2$$

c. **Calculate the concentration of each ion remaining in solution after precipitation is complete.**

To answer this question, think about what each ion is doing as a part of this reaction. There are a couple important things to keep in mind – is the ion part of the limiting reactant, and separately, is the ion going into forming the precipitate (effectively reducing the amount flowing around in solution as the free ion)?

Consider $OH^-$ first. $OH^-$ is part of the limiting reactant, so it is all used up, AND it is all going into forming the precipitate, a solid. Thus, there are no $OH^-$ ions floating around in solution. $[OH^-] = 0$

Consider $K^+$ next. Once again, $K^+$ is part of the limiting reactant, so it is all used up, BUT it is not doing anything chemically interesting from our perspective. $K^+$ is a spectator ion, and just stays floating in solution. So the amount that we started off with will be the amount we end up in solution with, in terms of moles. Thus,

$$0.02\ mol\ KOH * \frac{1\ mol\ K^+}{1\ mol\ KOH} = 0.02\ mol\ K^+$$

$$[K^+] = \frac{0.02\ mol\ K^+}{0.200\ L} 0.100\ M$$

Remember to use the combined volume when calculating molarity!

Now, let's consider $Mg^{2+}$ - it is not part of the limiting reactant, and some of it is going into the precipitate. However, some will be left over floating around in solution. We have to determine how much this is

$$0.02\ mol\ KOH * \frac{1\ mol\ Mg(NO_3)_2}{2\ mol\ KOH} * \frac{1\ mol\ Mg^{2+}}{1\ mol\ Mg(NO_3)_2} = 0.01\ mol\ Mg^{2+}$$

This indicates that since KOH is the limiting reactant, 0.01 mol of $Mg^{2+}$ is used up, going into the precipitate. Thus,

$$mol\ Mg^{2+} remaining = initial\ mol\ Mg^{2+} - used\ mol\ Mg^{2+}$$

$$\left(0.02\ mol\ Mg(NO_3)_2 * \frac{1\ mol\ Mg^{2+}}{1\ mol\ Mg(NO_3)_2}\right) - 0.01\ mol\ Mg^{2+} = 0.01\ mol\ Mg^{2+} remaining$$

$$[Mg^{2+}] = \frac{0.01\ mol\ Mg^{2+}}{0.200\ L} = 0.050\ M$$

Finally, $NO_3^-$. Once again, $NO_3^-$ is not going into forming the precipitate, and is not part of the limiting reactant. So, however much we started off with will be the amount that we end up with, floating around in solution.

$$0.02\ mol\ Mg(NO_3)_2 * \frac{2\ mol\ NO_3^-}{1\ mol\ Mg(NO_3)_2} = 0.04\ mol\ NO_3^-$$

$$[NO_3^-] = \frac{0.04\ mol\ NO_3^-}{0.2\ L} = 0.200\ M$$

8. **A solution is prepared by dissolving 0.5842 g of oxalic acid ($H_2C_2O_4$) in enough water to make 100.0 mL of solution. A 10.00 mL portion of this solution is then diluted to a final volume of 250.0 mL. What is the final molarity of the oxalic acid solution?**

$$0.5842\ g\ H_2C_2O_4 * \frac{1\ mol\ H_2C_2O_4}{90.03\ g\ H_2C_2O_4} = 0.00649\ mol\ H_2C_2O_4$$

$$[H_2C_2O_4] = \frac{0.00649\ mol\ H_2C_2O_4}{0.100\ L} = 0.0649\ M$$

Determine how many moles are removed in the 10.00 mL sample:

$$0.0649\ M = \frac{0.0649\ mol\ H_2C_2O_4}{1\ L} * 0.01\ L = 0.000649\ mol\ H_2C_2O_4$$

Calculate the final molarity:

$$[H_2C_2O_4] = \frac{0.000649\ mol\ H_2C_2O_4}{0.250\ L} = 2.596 * 10^{-3}\ M$$

Note that 10.00 mL was NOT added to 250 mL because the problem stated that 250 mL was the final volume of the solution.