General Chemistry I
PLI #6 AK
Thursday,10June 2020

1. **Vanillin, the dominant flavoring in vanilla, contains C, H, and O. When 1.05 g of thus substance is completely combusted, 2.43 g of CO_2 and 0.50 g of H_2O are produced. What is the empirical formula of vanillin?**

   Yet another combustion analysis problem!

$$0.50\ g\ H_2O * \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} * \frac{2\ mol\ H}{1\ mol\ H_2O} = 0.05549\ mol\ H * \frac{1.008\ g\ H}{1\ mol\ H} = 0.0559\ g\ H$$

$$2.43\ g\ CO_2 * \frac{1\ mol\ CO_2}{44.01\ g\ CO_2} * \frac{1\ mol\ C}{1\ mol\ CO_2} = 0.05521\ mol\ C * \frac{12.011\ g\ C}{1\ mol\ C} = 0.6632\ g\ C$$

$$Mass\ O = Mass\ sample - Mass\ C - Mass\ H = 1.05\ g - 0.6632\ g - 0.0559\ g$$
$$= 0.3309\ g\ O$$

$$0.3309\ g\ O * \frac{1\ mol\ O}{15.9994\ g\ O} = 0.0207\ mol\ O$$

   Take mole ratios to find empirical formula!

$$C: \frac{0.05521}{0.0207} = 2.67$$

$$H: \frac{0.05549}{0.0207} = 2.67$$

$$O: \frac{0.0207}{0.0207} = 1$$

   2.67 is equal to 8/3. Thus, multiply everything by 3 to get the empirical formula: $C_8H_8O_3$

2. **A mixture of $N_2(g)$ and $H_2(g)$ reacts in a closed container to form ammonia, $NH_3(g)$. The reaction ceases before either reactant has been totally consumed. At this stage, 3.0 mol N_2, 3.0 mol H_2, and 3.0 mol NH_3 are present. How many moles of N_2 and H_2 were present originally?**

   First write the balanced equation!

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Since we assume that we started off with zero product (zero ammonia), let's use that information to find out how much reactant was needed to generate that much product.

$$3.0\ mol\ NH_3 * \frac{1\ mol\ N_2}{2\ mol\ NH_3} = 1.5\ mol\ N_2\ needed$$

$$3.0\ mol\ NH_3 * \frac{3\ mol\ H_2}{2\ mol\ NH_3} = 4.5\ mol\ H_2\ needed$$

Thus, 1.5 mol N_2 and 4.5 mol H_2 were needed to bring the reaction to its current state, with 3 mol NH_3. However, there are also current mole amounts of each of the reactants, as noted in the problem. Thus,

$$Original\ N_2 amount = 1.5\ mol + 3.0\ mol = 4.5\ mol\ N_2$$

$$Original\ H_2 amount = 4.5\ mol + 3.0\ mol = 7.5\ mol\ H_2$$

3. **A mixture containing $KClO_3$, K_2CO_3, $KHCO_3$, and KCl was heated, producing CO_2, O_2, and H_2O gases according to the following equations:**

$$\boldsymbol{2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)}$$

$$\boldsymbol{2KHCO_3(s) \rightarrow K_2O(s) + H_2O(g) + 2CO_2(g)}$$

$$\boldsymbol{K_2CO_3(s) \rightarrow K_2O(s) + CO_2(g)}$$

The KCl does not react under the conditions of the reaction. If 100.0 g of the mixture produces 1.80 g of H_2O, 13.20 g of CO_2, and 4.00 g of O_2, what was the composition of the original mixture? (Assume complete decomposition of the mixture.)

This is similar to the combustion analysis in some ways. Recognize that all of the O_2 produced is due to the first reaction, and all of the H_2O produced comes from the second reaction.

$$4.00\ g\ O_2 * \frac{1\ mol\ O_2}{31.9988\ g\ O_2} * \frac{2\ mol\ KClO_3}{3\ mol\ O_2} * \frac{122.55\ g\ KClO_3}{1\ mol\ KClO_3} = 10.21\ g\ KClO_3$$

$$1.80\ g\ H_2O * \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} * \frac{2\ mol\ KHCO_3}{1\ mol\ H_2O} * \frac{100.12\ g\ KHCO_3}{1\ mol\ KHCO_3} = 20.00\ g\ KHCO_3$$

Next, we know how much CO_2 is produced, but that is produced in both the 2^{nd} and 3^{rd} reactions. So, since we just determined how much $KHCO_3$ was present, determine how much CO_2 that would have produced.

$$1.80\ g\ H_2O * \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} * \frac{2\ mol\ CO_2}{1\ mol\ H_2O} * \frac{44.01\ g\ CO_2}{1\ mol\ CO_2} = 8.79\ g\ CO_2$$

Thus, the CO_2 produced from reaction 3 is:

$$13.20\ g\ CO_2 - 8.79\ g\ CO_2 = 4.41\ g\ CO_2$$

Then, this can be used to find how much K_2CO_3 was present in the original mixture.

$$4.41\ g\ CO_2 * \frac{1\ mol\ CO_2}{44.01\ g\ CO_2} * \frac{1\ mol\ K_2CO_3}{1\ mol\ CO_2} * \frac{138.21\ g\ K_2CO_3}{1\ mol\ K_2CO_3} = 13.84\ g\ K_2CO_3$$

$$\begin{aligned} Mass\ KCl &= Mass\ sample - Mass\ KClO_3 - Mass\ KHCO_3 - Mass\ K_2CO_3 \\ &= 100.0 - 10.21 - 20.00 - 13.84 = 56.0\ g\ KCl \end{aligned}$$

Composition:

$$\%\ KCl = \frac{56.0\ g}{100.0\ g} = 56.0\%\ KCl$$

$$\%\ KClO_3 = \frac{10.21\ g}{100.0\ g} = 10.2\%\ KClO_3$$

$$\%\ KHCO_3 = \frac{20.00\ g}{100.0\ g} = 20.0\%\ KClO_3$$

$$\%\ K_2CO_3 = \frac{13.84\ g}{100.0\ g} = 13.8\%\ K_2CO_3$$

4. **Give the oxidation number of each element in the following compounds:**

   a. **BrO_3^-:** O: -2; Br: +5
   b. **H_2SO_4:** O: -2; S: +6; H: +1
   c. **CrO_4^{2-}:** O: -2; Cr: +6
   d. **$LiAlH_4$**: Li: +1; Al: +3; H: -1

5. **Your friend has heard that she can make ethanol by reacting C_2H_4 with H_2O under acidic conditions, but she's not sure how much of each starting material she needs. So she randomly mixes 101.7 g of C_2H_4 with 55.19 g of H_2O.**

$$\boldsymbol{C_2H_4 + H_2O \rightarrow C_2H_6O}$$

   a. **What is the theoretical yield of ethanol in mL (ethanol density = 0.789 g/mL)?**

$$101.7\ g\ C_2H_4 * \frac{1\ mol\ C_2H_4}{28.05\ g\ C_2H_4} * \frac{1\ mol\ C_2H_6O}{1\ mol\ C_2H_4} = 3.626\ mol\ C_2H_6O$$

$$55.19\ g\ H_2O * \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} * \frac{1\ mol\ C_2H_6O}{1\ mol\ H_2O} = 3.063\ mol\ C_2H_6O$$

Thus, H_2O is the limiting reactant

$$3.063\ mol\ C_2H_6O * \frac{46.07\ g\ C_2H_6O}{1\ mol\ C_2H_6O} * \frac{1\ mL}{0.789\ g} = 178.8\ mL\ C_2H_6O$$

b. How much (mass) excess reactant remains?

$$55.19\ g\ H_2O * \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} * \frac{1\ mol\ C_2H_4}{1\ mol\ H_2O} = 3.063\ mol\ C_2H_4\ used\ up$$

$$101.7\ g\ C_2H_4 * \frac{1\ mol\ C_2H_4}{28.05\ g\ C_2H_4} = 3.626\ mol\ C_2H_4\ initially$$

$$3.626\ mol - 3.063\ mol = 0.563\ mol\ C_2H_4\ remaining * \frac{28.05\ g}{1\ mol\ C_2H_4} = 15.8\ g\ C_2H_4\ remains$$

6. **Complete and balance the following molecular equations, and then write the net ionic equation for each (note in past answer keys, we have written H^+ in net ionic equations, but it is more correct to write H_3O^+ instead. H^+ doesn't actually exist itself in solution):**

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

$$HBr(aq) + NH_3(aq) \rightarrow Br^-(aq) + NH_4^+(aq)$$

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

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

$$H_3O^+(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + H_2O(l)$$

   b. **Aqueous hydrochloric acid and sodium acetate**

Overall balanced:

$$HCl(aq) + NaC_2H_3O_2(aq) \rightarrow HC_2H_3O_2(aq) + NaCl(aq)$$

Total ionic:

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + C_2H_3O_2^-(aq)$$
$$\rightarrow HC_2H_3O_2(aq) + Na^+(aq) + Cl^-(aq)$$

Net ionic:

$$H^+(aq) + C_2H_3O_2^-(aq) \rightarrow HC_2H_3O_2(aq)$$

$$H_3O^+(aq) + C_2H_3O_2^-(aq) \rightarrow HC_2H_3O_2(aq) + H_2O(l)$$

c. Aqueous perchloric acid and aqueous strontium hydroxide

Overall balanced:

$$2HClO_4(aq) + Sr(OH)_2(aq) \rightarrow Sr(ClO_4)_2(aq) + 2H_2O(l)$$

Total ionic:

$$2H^+(aq) + 2ClO_4^-(aq) + Sr^{2+}(aq) + 2OH^-(aq) \rightarrow Sr^{2+}(aq) + 2ClO_4^{2-}(aq) + 2H_2O(l)$$

Net ionic:

$$2H^+(aq) + 2OH^-(aq) \rightarrow 2H_2O(l)$$

$$2H_3O^+(aq) + 2OH^-(aq) \rightarrow 4H_2O(l)$$

The above equation can be simplified to:

$$H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(l)$$

7. Starting with solid sucrose, $C_{12}H_{22}O_{11}$, describe how you would

a. Prepare 250 mL of a 0.250 M sucrose solution

$$0.250\ M = \frac{0.250\ mol\ C_{12}H_{22}O_{11}}{1\ L} * 0.250\ L = 0.0625\ mol\ C_{12}H_{22}O_{11}$$

$$0.0625\ mol\ C_{12}H_{22}O_{11} * \frac{342.3\ g\ C_{12}H_{22}O_{11}}{1\ mol\ C_{12}H_{22}O_{11}} = 21.39\ g\ C_{12}H_{22}O_{11}$$

Measure out 21.39 g of sucrose, and add enough water to make 250 mL of solution.

b. Prepare 350.0 mL of 0.100 M $C_{12}H_{22}O_{11}$ starting with 3.00 L of 1.50 M $C_{12}H_{22}O_{11}$.

Determine how many moles of sucrose you need in the desired solution.

$$0.100\ M\ C_{12}H_{22}O_{11} = \frac{0.100\ mol\ C_{12}H_{22}O_{11}}{1\ L} * 0.350\ L = 0.035\ mol\ C_{12}H_{22}O_{11}$$

You should check to make sure that you have at least this many moles of sucrose to begin with. Otherwise, you wouldn't have enough to prepare the desired solution.

$$1.50\ M\ C_{12}H_{22}O_{11} = \frac{1.50\ mol\ C_{12}H_{22}O_{11}}{1\ L} * 3.00\ L = 4.50\ mol\ C_{12}H_{22}O_{11}$$

This confirms that we have enough moles of sucrose to make the desired solution. We need 0.035 mol of $C_{12}H_{22}O_{11}$.

$$0.035\ mol\ C_{12}H_{22}O_{11} * \frac{1\ L}{1.50\ mol\ C_{12}H_{22}O_{11}} = 0.0233\ L$$

So, we need 23.3 mL of the 1.50 M solution, and then the rest of the 350.0 mL volume should be water. The final solution volume is 350.0 mL.