CHEM 103
R&R 5
5 June 2024
Adapted from a 9 June 2021 document

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 quinone if you find that 0.105 g of quinone gives 0.257 g CO_2 and 0.0350 g H_2O when burned completely? Given a molecular weight of approximately 108 g mol^{-1}, what is its molecular formula?

$$0.257\,g\ CO_2 \cdot \frac{mol\ C}{44.01\,g\ CO_2} = 0.00584\ mol\ C \Longleftrightarrow 0.0701\ g\ C$$

$$0.0350\,g\ H_2O \cdot \frac{2\ mol\ H}{18.02\,g\ H_2O} = 0.00388\ mol\ H \Longleftrightarrow 0.00392\ g\ H$$

$$\underbrace{0.00194\ mol\ O}_{3:2:1\ \text{ratio}} \Longleftrightarrow 0.0310\ g\ O$$

$$54.05\,\frac{g}{mol} = \frac{1}{2}\ \text{molar mass}$$

emp. form. is C_3H_2O
molecular formula is $C_6H_4O_2$

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:

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

3. Balance the following equations:

a. $NH_4NO_2 \rightarrow N_2 + H_2O$

$NH_4NO_2 \longrightarrow N_2 + 2H_2O$

b. $P_4O_{10} + H_2O \rightarrow H_3PO_4$

$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$

c. $HCl + CaCO_3 \rightarrow CaCl_2 + CO_2 + H_2O$

$2HCl + CaCO_3 \longrightarrow CaCl_2 + CO_2 + H_2O$

d. $C_6H_{14} + O_2 \rightarrow CO_2 + H_2O$

$C_6H_{14} + \frac{19}{2}O_2 \longrightarrow 6CO_2 + 7H_2O$

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

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

total ionic is same.

net ionic is same.

b. Aqueous ammonia reacts with aqueous hydrochloric acid

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: $NH_3(aq) + H^+(aq) \longrightarrow NH_4^+(aq)$

c. Aqueous solutions of aluminum nitrate and potassium phosphate

balanced: $Al(NO_3)_3(aq) + K_3PO_4(aq) \longrightarrow AlPO_4(s) + 3KNO_3(aq)$

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

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 3$\underline{0}$0 g ethane with 650 g chlorine produced 490 g C_2H_5Cl.

$$3\underline{0}0 \text{ g } C_2H_6 \cdot \frac{\text{mol } C_2H_6}{30.068 \text{ g } C_2H_6} = 9.\underline{9}8 \text{ mol } C_2H_6$$

$$650 \text{ g } Cl_2 \cdot \frac{\text{mol } Cl_2}{70.906 \text{ g } Cl_2} = 9.\underline{1}7 \text{ mol } Cl_2 \leftarrow \text{limiting reactant}$$

$$490 \; C_2H_5Cl \cdot \frac{\text{mol } C_2H_5Cl}{64.513 \text{ g } C_2H_5Cl} = 7.\underline{5}95 \text{ mol } C_2H_5Cl$$

$$\text{yield} = \frac{7.\underline{5}95 \text{ mol}}{9.\underline{1}7 \text{ mol}} = 83\%$$

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.

a. Calculate the concentration of the overall solution, ammonium ions, and sulfate ions.

$$\text{molar mass } (NH_4)_2SO_4 = 2\left(14.01 \frac{g}{mol}\right) + 8\left(1.008 \frac{g}{mol}\right) + 32.07 \frac{g}{mol} + 4\left(16.00 \frac{g}{mol}\right) = 132.15 \text{ g/mol}$$

$$10.8 \text{ g } (NH_4)_2SO_4 \cdot \frac{\text{mol}}{132.15 \text{ g}} \cdot \frac{1}{0.1000 \text{ L}} = 0.817 \text{ M } (NH_4)_2SO_4 \leftarrow \text{STOCK SOLUTION}$$

Adding 50.00 mL water to 10.00 mL stock dilutes by factor of 6.

$\Longrightarrow$ Final $[(NH_4)_2SO_4] = 0.136$ M

$[SO_4^{2-}] = [(NH_4)_2SO_4] = 0.136$ M

$[NH_4^+] = 2[(NH_4)_2SO_4] = 0.272$ M

Explanation:

$M_1V_1 = M_2V_2$; $V_2 = V_1 + 5V_1 = 6V_1$

$M_1 \cancel{V_1} = M_2(6\cancel{V_1}) \Longrightarrow M_2 = \frac{1}{6} M_1$

b. In the above scenario, would your answers be the same if the problem instead had 10.8 g ammonium sulfate dissolved in 100.0 mL water? Why or why not?

No. Adding 100.0 mL water to the $(NH_4)_2SO_4$, which takes up space itself, would result in a volume slightly greater than 100.0 mL. As a result, the stock solution would have a slightly lower concentration.

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.

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

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

net: $Mg^{2+}(aq) + 2OH^-(aq) \longrightarrow Mg(OH)_2(s)$

b. What mass of precipitate forms?

$(.1000 L)(.200 M\ KOH) \cdot \frac{mol\ OH^-}{mol\ KOH} = 0.0200\ mol\ OH^-$] ← limiting reactant

$(.1000 L)(.200 M\ Mg(NO_3)_2) \cdot \frac{mol\ Mg^{2+}}{mol\ Mg(NO_3)_2} = 0.0200\ mol\ Mg^{2+}$

$0.0200\ mol\ OH^- \cdot \frac{mol\ Mg(OH)_2}{2\ mol\ OH^-} \cdot \frac{58.32\ g}{mol} = 0.583\ g\ Mg(OH)_2$

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

All OH^- used $\Rightarrow [OH^-] = 0$ M. K^+, NO_3^- diluted by factor of 2, with original $[NO_3^-] = 0.400$ M

$[K^+] = 0.100$ M ; $[NO_3^-] = 0.200$ M.

Half of Mg^{2+} used, then diluted by factor of 2 $\Rightarrow [Mg^{2+}] = 0.0500$ M

8. A solution is prepared by dissolving 0.5842 g 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?

molar mass: $2\left(1.008 \frac{g}{mol}\right) + 2\left(12.01 \frac{g}{mol}\right) + 4\left(16.00 \frac{g}{mol}\right) = 90.04 \frac{g}{mol}$

$0.5842\ g\ H_2C_2O_4 \cdot \frac{mol}{90.04\ g} \cdot \frac{1}{0.1000\ L} = 0.06488$ M ← orig. solution

$\underbrace{\text{Diluted by factor of 25}}_{\text{Explanation:}} \Longrightarrow [H_2C_2O_4] = \frac{1}{25}(0.06488\ M)$

$\boxed{= 2.595 \times 10^{-3}\ M}$

$M_1\cancel{V_1} = M_2V_2 = M_2(25\cancel{V_1})$

$M_1 = 25M_2 \Longrightarrow M_2 = \frac{1}{25}M_1$