HW 3-3 – Combustion Analysis; Molecular Formula **Name** Answer Key
Problems pg. 106 #3.98, 3.137, 3.52, 3.54, 3.119 and 2 extra problems

A. After combustion with excess oxygen, a 12.501 g sample of a petroleum compound produced 38.196 g of carbon dioxide and 18.752 of water. A previous analysis determined that the compound does not contain oxygen. Establish the empirical formula of the compound.

$$\left.\begin{array}{l} 38.196 \text{ g } CO_2 \times \dfrac{1 \text{ mol C}}{44.011 \text{ g } CO_2} = 0.86787 \text{ mol C} \div 0.86787 = 1 \text{ mol C} \times 5 = 5 \text{ mol C} \\ 18.752 \text{ g } H_2O \times \dfrac{2 \text{ mol H}}{18.016 \text{ g } H_2O} = 1.0817 \text{ mol H} \div 0.86787 = 2.3996 \text{ mol H} \times 5 = 11.998 \approx 12 \text{ mol H} \end{array}\right\} \boxed{C_5H_{12}}$$

3.98 A sample of a compound of Cl and O reacts with an excess of H_2 to give 0.233 g of HCl and 0.403 g of H_2O. Determine the empirical formula of the compound.
We assume that all the Cl in the compound ends up as HCl and all the O ends up as H_2O.

$$\left.\begin{array}{l} 0.233 \text{ g HCl} \times \dfrac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \times \dfrac{1 \text{ mol Cl}}{1 \text{ mol HCl}} = 0.00639 \text{ mol Cl} \div 0.00639 = 1 \times 2 = 2 \text{ Cl} \\ 0.403 \text{ g } H_2O \times \dfrac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \times \dfrac{1 \text{ mol O}}{1 \text{ mol } H_2O} = 0.0224 \text{ mol O} \div 0.00639 = 3.5 \times 2 = 7 \text{ O} \end{array}\right\} \boxed{Cl_2O_7}$$

3.137 When 12.1 g of *tert*-butyl ether (a compound of C, H, and O) are burned in an apparatus like the one shown in Fig 3.6, 30.1 g of CO_2 and 14.8 g of H_2O are formed. What is the empirical formula of the compound?

$$? \text{ g C} = 30.2 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 8.24 \text{ g C}$$

$$? \text{ g H} = 14.8 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 1.66 \text{ g H}$$

$$\text{mass O} = 12.1 \text{ g compound} - (8.24 \text{ g C} + 1.66 \text{ g H}) = 2.2 \text{ g O}$$

$$? \text{ mol C} = 8.24 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.686 \text{ mol C} \div 0.14 = 5$$

$$? \text{ mol H} = 1.66 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 1.65 \text{ mol H} \div 0.14 = 12$$

$$? \text{ mol O} = 2.2 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.14 \text{ mol O} \div 0.14 = 1$$

The empirical formula is **$C_5H_{12}O$**.

B. After combustion with excess oxygen, a 28.39 g sample of propanethiol produced 49.22 g of carbon dioxide, 26.87 of water and 29.84 g SO_3. The compound produced no other oxides. Determine the empirical formula of the compound.

$$\left.\begin{array}{l} 49.22 \text{ g } CO_2 \times \dfrac{1 \text{ mol C}}{44.01 \text{ g } CO_2} = 1.118 \text{ mol C} \div 0.3728 = 3 \text{ mol C} \\ 26.87 \text{ g } H_2O \times \dfrac{2 \text{ mol H}}{18.02 \text{ g } H_2O} = 2.982 \text{ mol H} \div 0.3728 = 8 \text{ mol H} \\ 29.848 \text{ g } SO_3 \times \dfrac{1 \text{ mol S}}{80.06 \text{ g } SO_3} = 0.3728 \text{ mol S} \div 0.3728 \text{ mol S} = 1 \end{array}\right\} \boxed{C_3H_8S}$$

3.52 The empirical formula of a compound is CH. If the molar mass of this compound is 78 g, what is its molecular formula?

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{78\text{ g}}{13.02\text{ g}} \approx 6$$, thus the molecular formula is $(CH)_6$, or $\mathbf{C_6H_6}$.

3.54 Monosodium glutamate (MSG) has the following composition by mass: 35.51% C, 4.77% H, 37.85% O, 8.29% N, and 13.60% Na. What is its molecular formula if its molar mass is 169 g?

$$n_C = 35.51\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} = 2.957\text{ mol C} \div 0.5916 = 4.998 \approx 5$$

$$n_H = 4.77\text{ g H} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = 4.73\text{ mol H} \div 0.5916 = 8$$

$$n_O = 37.85\text{ g O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 2.366\text{ mol O} \div 0.5916 = 3.999 \approx 4$$

$$n_N = 8.29\text{ g N} \times \frac{1\text{ mol N}}{14.01\text{ g N}} = 0.592\text{ mol N} \div 0.5916 = 1$$

$$n_{Na} = 13.60\text{ g Na} \times \frac{1\text{ mol Na}}{22.99\text{ g Na}} = 0.5916\text{ mol Na} \div 0.5916 = 1$$

Thus, the empirical formula is $\mathbf{C_5H_8O_4NNa}$.

Divide to find the factor for the molecular formula: $\dfrac{\text{molar mass}}{\text{empirical molar mass}} = \dfrac{169\text{ g}}{169.11\text{ g}} \approx 1$

Hence, the molecular formula and the empirical formula are the same, $\mathbf{C_5H_8O_4NNa}$.

3.119 Lysine contains C, H, O, and N. In one experiment, the complete combustion of 2.175 g of lysine gave 3.94 g CO_2 and 1.89 g H_2O. When reduced in hydrogen, a second 2.175 g sample of lysine produced 0.506 g NH_3. First determine the empirical formula for lysine, then given the molecular mass of lysine is 150 g/mol, determine its molecular formula.

$$?\text{ g C} = 3.94\text{ g CO}_2 \times \frac{1\text{ mol CO}_2}{44.01\text{ g CO}_2} \times \frac{1\text{ mol C}}{1\text{ mol CO}_2} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = 1.075\text{ g C}$$

$$?\text{ g H} = 1.89\text{ g H}_2\text{O} \times \frac{1\text{ mol H}_2\text{O}}{18.02\text{ g H}_2\text{O}} \times \frac{2\text{ mol H}}{1\text{ mol H}_2\text{O}} \times \frac{1.008\text{ g H}}{1\text{ mol H}} = 0.2114\text{ g H}$$

$$?\text{ g N} = 0.506\text{ g NH}_3 \times \frac{1\text{ mol NH}_3}{17.03\text{ g NH}_3} \times \frac{1\text{ mol N}}{1\text{ mol NH}_3} \times \frac{14.01\text{ g N}}{1\text{ mol N}} = 0.4165\text{ g N}$$

$$?\text{ g O} = 2.175\text{ g} - (1.075\text{ g} + 0.2114\text{ g} + 0.4165\text{ g}) = 0.4721\text{ g O}$$

$$?\text{ mol C} = 1.075\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} = 0.08951\text{ mol C} \div 0.02951 = 3$$

$$?\text{ mol H} = 0.2114\text{ g H} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = 0.2097\text{ mol H} \div 0.02951 = 7$$

$$?\text{ mol N} = 0.4165\text{ g N} \times \frac{1\text{ mol N}}{14.01\text{ g N}} = 0.02973\text{ mol N} \div 0.02951 = 1$$

$$?\text{ mol O} = 0.4721\text{ g O} \times \frac{1\text{ mol O}}{16.00\text{ g O}} = 0.02951\text{ mol O} \div 0.02951 = 1$$

Thus, the empirical formula is $\mathbf{C_3H_7NO}$.

$\dfrac{150\ g}{73.10\ g} \approx 2 \Rightarrow$ molecular formula is $(C_3H_7NO)_2$ or $\mathbf{C_6H_{14}N_2O_2}$.