Problem 2

Problem 2. A second compound (also containing Carbon, Nitrogen, Oxygen, and Hydrogen) is analyzed in the same way as the above. When a 10.0 gram sample of this compound is combusted, and it produces 15.17 grams of carbon dioxide, 4.69 grams of water, and 3.96 grams of nitrogen dioxide. The molecular weight of this compound is between 200 and 250 g/mol.

a. What is the empirical formula of this compound?
b. What is the molecular formula of this compound?

$$C_wH_xN_yO_z + O_2 \rightarrow CO_2 + H_2O + NO_2$$

10.0 g 15.17g 4.69g 3.96g

a)

$$(15.17g\ CO_2)\left(\frac{12.011g\ C}{44.0098g\ CO_2}\right) = 4.1401 \text{ grams carbon}$$

$$(4.69\ g\ H_2O)\left(\frac{2\times 1.0079\ g\ H}{18.0152g\ H_2O}\right) = 0.52478 \text{ grams hydrogen}$$

$$(3.96\ g\ NO_2)\left(\frac{14.0067g\ N}{46.0055g\ NO_2}\right) = 1.20565 \text{ grams nitrogen}$$

$$10.0\ g\ C_wH_xN_yO_z - \underset{\text{carbon}}{4.1401g} - \underset{\text{hydrogen}}{0.52478g} - \underset{\text{nitrogen}}{1.20565g} = \underset{\text{oxygen}}{4.1295g}$$

$$(4.1401g\ C)\left(\frac{1\text{mole}}{12.011g}\right) = 0.3447 \text{ moles C}$$

$$(.52478g\ H)\left(\frac{1\text{mole}}{1.0079\ g}\right) = 0.52067 \text{ moles H}$$

$$(1.20565g\ N)\left(\frac{1\text{mole}}{14.0067g}\right) = 0.08608 \text{ moles N}$$

$$\frac{4.1295g\ O}{15.9994\ g/\text{mole}} = 0.25810 \text{ moles O}$$

$$\frac{.3447}{.08608} = 4.004 \approx 4$$

$$\frac{.52067}{.08608} = 6.049 \approx 6$$

$$\frac{.08608}{.08608} = 1$$

$$\frac{.25810}{.08608} = 2.998 \approx 3$$

So, empirical formula is $\boxed{C_4H_6NO_3}$ (a)

empirical mass of $C_4H_6NO_3$ is 116 amu

116 × 2 = 232   232 is between 200 and 250 amu

(b)

So molecular formula is $(C_4H_6NO_3)2 = \boxed{C_8H_{12}N_2O_6}$

molecular formula

Problem 3

Problem 3. (from the 2000 AP test) Answer the following questions about BeC_2O_4 and its hydrate.

a. Calculate the mass percent of carbon in the hydrated form of the solid that has the formula $BeC_2O_4 \cdot 3H_2O$.

b. When heated to 220. °C, $BeC_2O_4 \cdot 3H_2O_{(s)}$ dehydrates completely as represented below.

$$BeC_2O_4 \cdot 3H_2O_{(s)} \text{ -----> } BeC_2O_{4(s)} + 3H_2O_{(g)}$$

If 3.21 grams of $BeC_2O_4 \cdot 3H_2O_{(s)}$ is heated to 220. °C, calculate

(i) the mass of $BeC_2O_{4(s)}$ formed, and

(ii) the volume of $H_2O_{(g)}$ released, measured at 220. °C and 735 mmHg.

(Note: gas laws will NOT be covered on the first unit test, though you'll need one for the above problem)

(a) $$\%\ \text{Carbon} = \frac{2(12.011)\ \text{amu}}{(2(12.011) + 9.0122 + 4(15.9994) + 3(18.0152))\ \text{amu}} \times 100$$

$$= \frac{2(12.011)}{151.0774} \times 100 = \boxed{15.90\,\%}\ \text{carbon.}$$

(b) (i) $$\left(3.21\ g\ BeC_2O_4 \cdot 3H_2O\right)\left(\frac{97.0318\ g\ BeC_2O_4}{151.0774\ g\ BeC_2O_4 \cdot 3H_2O}\right) = \boxed{2.06\ g\ BeC_2O_{4(s)}}$$

or $$\left(3.21\ g\ BeC_2O_4 \cdot 3H_2O\right)\left(\frac{1\ \text{mole}\ BeC_2O_4 \cdot 3H_2O}{151.0774\ g}\right)\left(\frac{1\ \text{mole}\ BeC_2O_4}{1\ \text{mole}\ BeC_2O_4 \cdot 3H_2O}\right)\left(\frac{97.0318\ g}{1\ \text{mole}\ BeC_2O_4}\right) =$$

(b) (ii) $$\left(3.21\ g\ BeC_2O_4 \cdot 3H_2O\right)\left(\frac{1\ \text{mole}}{151.0774\ g}\right)\left(\frac{3\ \text{mole}\ H_2O}{1\ \text{mole}\ BeC_2O_4 \cdot 3H_2O}\right) = .0637(42)\ \text{moles}\ H_2O$$

$$PV = nRT$$

$$V = \frac{nRT}{P} = \frac{(.0637(42)\ \text{moles})(.0821\ \frac{l \cdot atm}{mol \cdot K})(220. + 273)\,K}{(735/760.)\ atm}$$

$$= \boxed{2.67\ \text{liters}\ H_2O_{(g)}}$$

PROBLEM 4

Problem 4. (from the 2005 AP test)

Answer the following questions about a pure compound that contains only carbon, hydrogen, and oxygen.

a. A 0.7549 g sample of the compound burns in $O_{2(g)}$ to produce 1.9061 g of $CO_{2(g)}$ and 0.3370 g of $H_2O_{(g)}$.
   (i) Calculate the individual masses of C, H, and O in the 0.7549 g sample.
   (ii) Determine the empirical formula for the compound.

b. (In part b, the test gave some freezing point depression data, and asked you to use this to solve for the molar mass of the unknown compound. The answer for the molar mass of the unknown compound, was 122 g/mole. Nothing to write here, but this may be helpful in answering part c, below.)

c. Without doing any calculations, explain how to determine the molecular formula of the compound based on the answers to parts (a) (ii) and (b).

(a) (i) $(1.9061\text{ g } CO_2)\left(\dfrac{12.011\text{ g C}}{44.0098\text{ g } CO_2}\right) = \boxed{0.5202 \text{ grams carbon}}$

$(0.3370\text{ g } H_2O)\left(\dfrac{2.0158\text{ g H}}{18.0152\text{ g } H_2O}\right) = \boxed{0.03771 \text{ grams hydrogen}}$

$0.7549\text{ g total} - 0.5202\text{ g C} - 0.03771\text{ g H} = \boxed{0.1970 \text{ grams oxygen}}$

(ii) $(.5202\text{ g C})\left(\dfrac{1\text{ mole}}{12.011\text{ g}}\right) = .04331 \text{ moles C}$

$(.03771\text{ g H})\left(\dfrac{1\text{ mole}}{1.0079\text{ g}}\right) = 0.03741 \text{ moles H}$

$(0.1970\text{ g O})\left(\dfrac{1\text{ mole}}{15.9994\text{ g}}\right) = 0.01231 \text{ moles O}$

$\dfrac{.04331}{.012131} = 3.517 \text{ moles C per mole O}$

$\dfrac{.03741\text{ mole H}}{.012131\text{ moles O}} = 3.084 \text{ moles H per mole O}$

$(C_{3.5}H_3O_1) \times 2 = \boxed{C_7H_6O_2}$ empirical formula

(c) To determine the molecular formula, add up the empirical mass of the compound, based on the empirical formula found in a(ii). Divide the molar mass of the compound (found in part b) by the empirical mass to find a whole number. Multiply all of the subscripts in the empirical formula by this whole number to find the molecular formula.

(Note: you don't actually have to do the math, though it is nice to check work by seeing that the empirical mass of 122 amu for $C_7H_6O_2$ matches the molar mass of 122 amu, so they are related by a whole number.....of one. So in this case, empirical formula = molecular formula)