KEY

GENERAL & ANALYTICAL CHEMISTRY I
CHMG-141
With Dr. Bailey

Name____________

Recitation
Week 9
Molecular and Empirical Formulas

Concepts you should know:

1) Finding % composition of elements in a compound.
2) Convert % composition to mole composition → to empirical formula → to molecular formula.

1) What is the mass % of F in the compound KrF_2?

$$\frac{2 \times (19.00)\,g}{\underbrace{83.80\,g + 2 \times (19.00)\,g}_{121.80}} \times 100\% = 31.20\%$$

2) Pyrophosphoric acid is made of 2.27% hydrogen and 34.80% phosphorus. The rest is oxygen. The molar mass of pyrophosphoric acid is 177.97 grams/mole.

Determine the <u>empirical and molecular formula</u> for pyrophosphoric acid.

Strategy:

a. Determine the mass of each component in a 100.0-gram sample.
b. How many moles of each element are in the sample?
c. Determine the smallest whole number ratio of the number of moles of each element.
d. Determine the <u>empirical formula</u> for pyrophosphoric acid.
e. Determine the molar mass of the empirical formula for pyrophosphoric acid.
f. Determine the <u>molecular formula</u>.

Pyrophosphoric acid is made of 2.27% hydrogen and 34.80% phosphorus. The rest is oxygen.

a) Determine the mass of each component in a 100.0-gram sample.

2.27% H by mass → 2.27 g H
34.80% P by mass → 34.80 g P
62.93 % O by mass → 62.93 g O

b) How many moles of each element are in the 100.0 gram sample?

2.27 g H x [1 mole H÷ 1.01 g H] = 2.25 mole
34.80 g P X [1 mole P÷ 30.97 g P] = 1.124 mole
62.93 g O X [1 mole O÷ 15.99 g O] = 3.934 mole

c) Determine the smallest whole number ratio of the number of moles of each element. This is the empirical formula for pyrophosphoric acid.

H = 2.248 ÷ 1.124 = 2 → 4
P = 1.124÷ 1.124 = 1 → 2
O = 3.934 ÷ 1.124 = 3.5 → 7
Acids usually have the H written first and the O last. $H_4P_2O_7$

d) Determine the mole mass of the empirical formula for pyrophosphoric acid.

$H_4P_2O_7$ Mole Mass = 4[1.01] + 2[30.97] + 7[15.99] = 177.91

e) The mole mass of pyrophosphoric acid is 177.97 grams/mole. Determine the molecular formula.

Empirical formula = molecular formula

f). $H_4P_2O_7$

3) A compound contains 12.0 grams of carbon, 3.00 grams of H and 8.00 grams of O.
Write the empirical and molecular formula for the compound

a. Convert the number of grams of each element into moles.
b. Convert the mole ratio into an <u>empirical formula</u>
c. If the molar mass is 92.0 grams, determine <u>the molecular formula.</u>

a) C $12.0g \times \frac{1 \text{ mol C}}{12.01 g} = 1.00$ mole

H $3.00 g \times \frac{1 \text{ mol H}}{1.01 g} = 3.00$ mole

O $8.00 g \times \frac{1 \text{ mole O}}{16.00 g} = 0.500$ mole

b) $C_{1.00} H_{3.00} O_{0.500}$

Double the subscripts $\boxed{C_2 H_6 O}$

Convert the mole ratio into an Empirical Formula

c). If the mole mass is 92.0 g, determine the molecular formula.

$MM\ C_2H_6O = 2\times(12.01) + 6\times(1.01) + 1\times(16.00) = 46 \frac{g}{mol}$

$\frac{92.0 g}{46.0 g} = 2$ = Twice the Empirical formula mass

$C_{2\times2} H_{6\times2} O_{1\times2} = \boxed{C_4 H_{12} O_2}$

molecular formula

4) An unknown compound contains only carbon, hydrogen, and oxygen ($C_x H_y O_z$). Combustion of 6.50 g of this compound produced 9.53 g of carbon dioxide and 3.90 g of water.

Write the empirical formula for the compound

a) How many moles of carbon, C, were in the original sample?

b) How many moles of hydrogen, H, were in the original sample?

c) How many moles of oxygen, O, were in the original sample?

d) Convert the mole ratio into an empirical formula

a) $$9.53\, g_{CO_2} \times \frac{1 \text{ mole } CO_2}{(12.01 + 2 \cdot 16.00)\, g_{CO_2}} \times \frac{1 \text{ mole } C}{1 \text{ mole } CO_2} = \boxed{0.2165 \text{ mole } C}$$

b) $$3.90\, g_{H_2O} \times \frac{1 \text{ mole } H_2O}{(2 \times 1.01 + 16.00)\, g_{H_2O}} \times \frac{2 \text{ mole } H}{1 \text{ mole } H_2O} = \boxed{0.4329 \text{ mole } H}$$

c). $$0.2165 \text{ mole}_C \times \frac{12.01 g}{1 \text{ mole } C} = 2.601\, g\, C$$

$$0.4329 \text{ mole}_H \times \frac{1.008 g}{1 \text{ mole } H} = 0.4364\, g\, H$$

$$6.50 g_{Total} - (2.601 g_C + 0.4364\, g_H) = 3.463 g_O$$

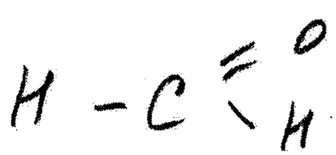
$$3.463\underline{3}\, g_O \times \frac{1\,\text{mole}\, O}{16.00\, g_O} = \boxed{0.216\underline{4}\,\text{mole}_O}$$

d). $C_{0.2165}\, H_{0.4329}\, O_{0.2164}$

$$C_{\frac{0.2165}{0.2164}}\; H_{\frac{0.4329}{0.2164}}\; O_{\frac{0.2164}{0.2164}}$$

$$C_1\, H_{1.999}\, C_1$$

$$\boxed{CH_2O}$$



This is formaldehyde