Practice Problems, solutions

1. Caffeine has the following percent composition: carbon 49.48%, hydrogen 5.19%, oxygen 16.48% and nitrogen 28.85%. What is its empirical formula?

Imagine that you have 100.00g of pure caffeine:

$$49.48 \text{ g C} \times \frac{1 \text{ mole C}}{12.01 \text{ g C}} = 4.120 \text{ mole C}$$

$$5.19 \text{ g H} \times \frac{1 \text{ mole H}}{1.008 \text{ g H}} = 5.149 \text{ mole H}$$

$$16.48 \text{ g C} \times \frac{1 \text{ mole O}}{16.00 \text{ g O}} = 1.030 \text{ mole O}$$

$$28.85 \text{ g N} \times \frac{1 \text{ mole N}}{14.01 \text{ g N}} = 2.059 \text{ mole N}$$

Divide each number of moles by the lowest number of moles to give the molar ratio:

$$\frac{4.120 \text{ mole C}}{1.030 \text{ mole O}} = \frac{4 \text{ C}}{1 \text{ O}}$$

$$\frac{5.149 \text{ mole H}}{1.030 \text{ mole O}} = \frac{5 \text{ H}}{1 \text{ O}}$$

$$\frac{2.059 \text{ mole N}}{1.030 \text{ mole O}} = \frac{2 \text{ N}}{1 \text{ O}}$$

So, the formula is $C_4H_5N_2O$

2. What is the mass percent of hydrogen in $C_{10}H_{14}N_2$? What is the empirical formula for this compound?

$$\%H = \frac{\text{mass of H}}{\text{mass of 1 mole of } C_{10}H_{14}N_2} = \frac{14 \times 1.008 \text{ g}}{162.23 \text{ g}} \times 100\% = 8.699\% \text{ H}$$

Empirical formula: C_5H_7N

3. If 1.951 g $BaCl_2 \cdot nH_2O$ yields 1.864 g of anhydrous $BaSO_4$ after treatment with sulfuric acid, calculate n. ("anhydrous" means without water)

In this question, you finding out the moles of barium is the first step, which you can do because you have a known amount of $BaSO_4$:

$$1.864 \text{ g } BaSO_4 \times \frac{1 \text{ mole } BaSO_4}{233.37 \text{ g } BaSO_4} \times \frac{1 \text{ mole Ba}}{1 \text{ mole } BaSO_4} = 0.007987 \text{ mole Ba}$$

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Next, consider what mass of barium that is, and what mass of chlorine you have as a result:

$$0.007987 \text{ mole Ba} \times \frac{137.3\text{g Ba}}{1 \text{ mole Ba}} = 1.097 \text{ g Ba}$$
$$0.007987 \text{ mole Ba} \times \frac{2 \text{ mole Cl}}{1 \text{ mole Ba}} \times \frac{35.45 \text{ g Cl}}{1 \text{ mole Cl}} = 0.05663 \text{ g Cl}$$

Determine the mass of water by difference:

$$1.951 \text{ g } BaCl_2 \bullet nH_2O - 1.097 \text{ g Ba} - 0.05663 \text{ g Cl} = 0.2877 \text{ g } H_2O$$

Finally, determine the number of moles of water and the mole ratio of water to barium:

$$0.2877 \text{ g } H_2O \times \frac{1 \text{ mole } H_2O}{18.02 \text{ g } H_2O} = 0.01597 \text{ mole } H_2O$$
$$n = \frac{0.01597 \text{ mole } H_2O}{0.007987 \text{ mole Ba}} = 2$$

4. Elemental analysis is performed on 1.500 g of a compound revealed that 0.8179 g is carbon, 0.5448 g is oxygen, and the remainder is hydrogen. What is the empirical formula of the compound?

$$0.8179 \text{ g C} \times \frac{1 \text{ mole C}}{12.01 \text{ g C}} = 0.06810 \text{ mole C}$$
$$0.5448 \text{ g O} \times \frac{1 \text{ mole O}}{16.00 \text{ g O}} = 0.03405 \text{ mole O}$$
$$0.1373 \text{ g H} \times \frac{1 \text{ mole H}}{1.008 \text{ g H}} = 0.1362 \text{ mole H}$$

Divide each number of moles by the lowest number of moles to give the molar ratio:

$$\frac{0.06810 \text{ mole C}}{0.03405 \text{ mole O}} = \frac{2\text{C}}{1\text{O}}$$
$$\frac{0.1362 \text{ mole H}}{0.03405 \text{ mole O}} = \frac{4\text{H}}{1\text{O}}$$

Empirical formula: C_2H_4O

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5. The molar mass of the compound in problem 4 is 132.16 g mol^{-1}. What is the molecular formula of the compound?

   Divide molar mass of the molecular formula by the molar mass of the empirical formula:

   $\frac{132.16}{44.05} = \frac{3}{1}$ *So, multiply the empirical formula by a factor of 3:* $C_6H_{12}O_3$

6. The reaction between hydrazine, N_2H_4, and hydrogen peroxide, H_2O_2, has been used as a rocket fuel, given as the following un-*balanced* equation:

$$N_2H_4(l) + 7H_2O_2(l) \rightarrow 2HNO_3(aq) + 8H_2O(l)$$

Balance the equation. According to this equation determine the following:

a) If 3.25 moles of N_2H_4 react, how many moles of HNO_3 are formed?

$$3.25 \text{ mole } N_2H_4 \times \frac{2 \text{ mole } HNO_3}{1 \text{ mole } N_2H_4} = 7.50 \text{ mole } HNO_3$$

b) If 10.12 moles of water are formed, how many moles of nitric acid are formed?

$$10.12 \text{ mole } H_2O \times \frac{2 \text{ mole } HNO_3}{8 \text{ mole } H_2O} = 2.53 \text{ mole } H_2O$$

7. In dilute nitric acid, copper metal dissolves according to the following un-*balanced* equation:

$$Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

   Balance the equation. According to this equation determine the following:

a) How many grams of nitric acid are required to dissolve 11.45 g of Cu?

$$11.45 \text{ g Cu} \times \frac{1 \text{ mole Cu}}{63.55 \text{ g Cu}} \times \frac{4 \text{ mole } HNO_3}{1 \text{ mole Cu}} \times \frac{45.42 \text{ g } HNO_3}{1 \text{ mole } HNO_3} = 45.42 \text{ g } HNO_3$$

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b) How many grams of NO_2 gas are formed when 11.45 grams of Cu react?

$$11.45 \text{ g Cu} \times \frac{1 \text{ mole Cu}}{63.55 \text{ g Cu}} \times \frac{2 \text{ mole } NO_2}{1 \text{ mole Cu}} \times \frac{46.01 \text{ g } NO_2}{1 \text{ mole } NO_2} = 16.58 \text{ g } NO_2$$

c) If 44.0 g of copper are allowed to react with 65.1 g of nitric acid, how many grams of copper(II) nitrate are formed?

$$44.0 \text{ g Cu} \times \frac{1 \text{ mole Cu}}{63.55 \text{ g Cu}} \times \frac{1 \text{ mole } Cu(NO_3)_2}{1 \text{ mole Cu}} \times \frac{187.57 \text{ g } Cu(NO_3)_2}{1 \text{ mole } Cu(NO_3)_2} = 130. \text{ g } Cu(NO_3)_2$$

$$65.1 \text{g } HNO_3 \times \frac{1 \text{ mole } HNO_3}{63.02 \text{ g } HNO_3} \times \frac{1 \text{ mole } Cu(NO_3)_2}{4 \text{ mole } HNO_3} \times \frac{187.57 \text{ g } Cu(NO_3)_2}{1 \text{ mole } Cu(NO_3)_2} = 48.4 \text{ g } Cu(NO_3)_2$$

So, 48.4 g of $Cu(NO_3)_2$ are formed and nitric acid is the limiting reactant.

d) In part c, if 34.6 g of copper (II) nitrate are actually formed, what is the percent yield?

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{34.6 \text{ g}}{48.4 \text{ g}} \times 100 = 71.4\%$$

e) In part c, which reactant is in excess, and how much of it is left over?

Cu is in excess. First, calculate how much Cu is consumed:

$$65.1 \text{g } HNO_3 \times \frac{1 \text{ mole } HNO_3}{63.02 \text{ g } HNO_3} \times \frac{1 \text{ mole Cu}}{4 \text{ mole } HNO_3} \times \frac{63.55 \text{ g Cu}}{1 \text{ mole Cu}} = 16.4 \text{ g Cu used}$$

Then determine the amount left over by difference: 44.0g Cu-16.4 g Cu = 27.6 g Cu remaining.

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8. Determine the percent yield for the reaction below if a mixture of 98.0 grams of HNO_3 and 119 grams of $Ca(OH)_2$ yielded 105 grams of calcium nitrate. Remember to balance the equation first.

$$Ca(OH)_2(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(l)$$

$$98.0 \text{ g } HNO_3 \times \frac{1 \text{ mole } HNO_3}{63.02 \text{ g } HNO_3} \times \frac{1 \text{ mole } Ca(NO_3)_2}{2 \text{ mole } HNO_3} \times \frac{164.1 \text{ g } Ca(NO_3)_2}{1 \text{ mole } Ca(NO_3)_2} = 127.6 \text{ g } Ca(NO_3)_2$$

$$119 \text{ g } Ca(OH)_2 \times \frac{1 \text{ mole } Ca(OH)_2}{74.10 \text{ g } Ca(OH)_2} \times \frac{1 \text{ mole } Ca(NO_3)_2}{1 \text{ mole } Ca(OH)_2} \times \frac{164.1 \text{ g } Ca(NO_3)_2}{1 \text{ mole } Ca(NO_3)_2} = 263.5 \text{ g } Ca(NO_3)_2$$

Thus, the theoretical yield of calcium nitrate is 127.6 g.

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{105 \text{ g}}{127.6 \text{ g}} \times 100 = 82.3\%$$

9. Aluminum reacts with hydrochloric acid to produce aluminum chloride and hydrogen gas, as shown in the unbalanced reaction below. How many liters of hydrogen gas could be produced from a bar of aluminum with dimensions of 3.05 cm x 2.01 cm x 40.8 cm? The density of hydrogen at reaction conditions is 0.0899 g/L and the density of the aluminum is 2.702 g/cm^3.

$$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$$

First, compute the volume of Al: 3.05 cm x 2.01 cm x 40.8 cm = 250.1 cm^3

$$250.1 \text{ cm}^3 \text{ Al} \times \frac{2.702 \text{ g Al}}{1 \text{ cm}^3} \times \frac{1 \text{ mole Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mole } H_2}{2 \text{ mole Al}} \times \frac{2.016 \text{ g } H_2}{1 \text{ mole } H_2} \times \frac{1 \text{ L } H_2}{0.0899 \text{ g } H_2} = 843 \text{ L } H_2$$

10. Combustion analysis of naphthalene produced 8.80g of CO_2 and 1.44 g of H_2O. Determine the empirical formula of naphthalene.

$$8.80\text{g } CO_2 \times \frac{1 \text{ mole } CO_2}{44.01\text{g } CO_2} \times \frac{1 \text{ mole C}}{1 \text{ mole } CO_2} = 0.200 \text{ mole C}$$

$$1.44\text{g } H_2O \times \frac{1 \text{ mole } H_2O}{18.02 \text{ g } H_2O} \times \frac{2 \text{ mole H}}{1 \text{ mole } H_2O} = 0.160 \text{ mole H}$$

$$\frac{0.200 \text{ mole C}}{0.160 \text{ mole H}} = \frac{1.25 \text{ mole C}}{1 \text{ mole H}} \times \frac{4}{4} = \frac{5 \text{ mole C}}{4 \text{ mole H}} \Rightarrow C_5H_4$$