# Stoichiometry Problems Worksheet – Answer Key

1. Assume a sample of 100 g, then there are

   12.17 g C or 1.0132 mol C
   0.51 g H or 0.5060 mol H
   40.48 g Br or 0.5067 mol Br
   17.96 g Cl or 0.5066 mol Cl
   28.87 g F or 1.5196 mole F

   The element with the smallest number of moles is hydrogen so we calculate the mole ratio of other elements to hydrogen. Upon doing so, we find that there is one bromine and one chlorine for each hydrogen, that there are two carbons for each hydrogen and that there are three fluorines for each hydrogen. The empirical formula, therefore, is $C_2HBrClF_3$. Given that there is one hydrogen in the molecule, then the molecular and empirical formulas are the same. The molar mass, therefore, is 197.38 g/mol.

2. Because the original compound must contain oxygen and the final compound contains only K and Cl, we know that the original compound must have an empirical formula of $K_xCl_yO_z$. Assuming 100 g of the original compound gives

   31.9 g K or 0.8159 mol K
   28.9 g Cl or 0.8152 mol Cl
   39.2 g O or 2.4501 mol O

   which is consistent with an empirical formula of $KClO_3$. For the product, a 100 g sample has

   52.4 g K or 1.340 mol K
   47.6 g Cl or 1.343 mol Cl

   and an empirical formula of KCl. The balanced reaction, therefore, is

   $$2KClO_3 \rightarrow 2KCl + 3O_2$$

3. Beginning with 5.00 g of $BaCl_2 \cdot xH_2O$ and ending with 4.26 g (0.02046 mol) of $BaCl_2$ means that 0.74 g (0.04108 mol) of $H_2O$ were removed. The mole ratio between $BaCl_2$ and $H_2O$ is 1:2; thus, the formula of the hydrate is $BaCl_2 \cdot 2H_2O$.

4. Tackle this with conversion factors, starting with the known ratio of 4 Fe atoms for every molecule of hemoglobin (or 4 mol Fe for every mol of hemoglobin); thus

   $$\frac{4 \text{ mol Fe}}{\text{mol hemoglobin}} \times \frac{55.847 \text{ g Fe}}{\text{mol Fe}} \times \frac{100 \text{ g hemoglobin}}{0.355 \text{ g Fe}} = 66{,}700 \text{ g/mol}$$

5. To solve this, just string together the stoichiometry.

$$31.5\ \text{mL S}_2\text{O}_3^{2-} \times \frac{1.00\ \text{mol S}_2\text{O}_3^{2-}}{\text{L}} \times \frac{1\ \text{L}}{1000\ \text{mL}} \times \frac{1\ \text{mol I}_3^-}{2\ \text{mol S}_2\text{O}_3^{2-}} \times$$

$$\frac{2\ \text{mol Cu}^{2+}}{\text{mol I}_3^-} \times \frac{1\ \text{mol Cu}}{1\ \text{mol Cu}^{2+}} \times \frac{63.546\ \text{g Cu}}{\text{mol Cu}} = 2.002\ \text{g Cu}$$

The percentage of Cu in the original sample, therefore, is

$$\frac{2.002\ \text{g Cu}}{2.50\ \text{g sample}} \times 100 = 80.1\%\ \text{w/w Cu}$$

6. If the product weighs 231.6 g and the original amount of iron is 167.6 g (3.00 moles), then the amount of oxygen in the iron oxide is 64.0 g or 4.00 moles. The mole ratio of 3:4 shows that the oxide is $Fe_3O_4$.

7. This is essentially a limiting reagent problem, with the solution's degree of red color determined by the molarity of $Br_2$ at the end of the reaction. Remembering that

$$\text{moles} = \text{M} \times \text{V}$$

we have the following:

For the first experiment, the reaction mixture contains $1.00\times10^{-3}$ moles of $Br^-$ and $1.00\times10^{-3}$ moles of $Cl_2$; bromide is the LR and $5.00\times10^{-4}$ moles of $Br_2$ are produced. The final concentration of $Br_2$ is

$$\frac{5.00\times10^{-4}\ \text{moles Br}_2}{0.150\ \text{L}} = 3.33\times10^{-3}\ \text{M Br}_2$$

For the second experiment, the reaction mixture contains $2.00\times10^{-3}$ moles of $Br^-$ and $5\times10^{-4}$ moles of $Cl_2$. This time, $Cl_2$ is the LR and $5.00\times10^{-4}$ moles of $Br_2$ are produced. The final concentration of $Br_2$ is

$$\frac{5.00\times10^{-4}\ \text{moles Br}_2}{0.150\ \text{L}} = 3.33\times10^{-3}\ \text{M Br}_2$$

Because the same concentration of $Br_2$ is the same in each experiment, the solutions will have the same shade of red (observation C).