Solutions to the Extra Problems for Chapter 8

1. The answer is 83.4%. To figure out percent yield, you first have to determine what stoichiometry says should be made:

$$\text{Mass of } MgCl_2 = 24.31 \text{ amu} + 2\times 35.45 \text{ amu} = 95.21 \text{ amu}$$

$$1 \text{ mole } MgCl_2 = 95.21 \text{ g } MgCl_2$$

$$\frac{50.0 \cancel{\text{ g } MgCl_2}}{1} \times \frac{1 \text{ mole } MgCl_2}{95.21 \cancel{\text{ g } MgCl_2}} = 0.525 \text{ moles } MgCl_2$$

$$\frac{0.525 \cancel{\text{ moles } MgCl_2}}{1} \times \frac{2 \text{ moles } AgCl}{1 \cancel{\text{ mole } MgCl_2}} = 1.05 \text{ moles } AgCl$$

$$\text{Mass of } AgCl = 107.87 \text{ amu} + 35.45 \text{ amu} = 143.32 \text{ amu}$$

$$1 \text{ mole } AgCl = 143.32 \text{ g } AgCl$$

$$\frac{1.05 \cancel{\text{ moles } AgCl}}{1} \times \frac{143.32 \text{ g } AgCl}{1 \cancel{\text{ mole } AgCl}} = 1.50 \times 10^2 \text{ g } AgCl$$

Because of the 1.05, we need three significant figures. Thus, we have to report this in scientific notation. The actual yield is what the chemist actually got (125.1 g), and the theoretical yield is what stoichiometry says he should have gotten (1.50×10^2 g):

$$\text{Percent Yield} = (\text{Actual Yield} \div \text{Theoretical Yield})\times 100\% = (125.1 \text{ g} \div 1.50\times10^2 \text{ g})\times 100\% = 83.4\%$$

2. The actual yield is 163 g. The percent yield can tell us what she actually got, but in order to use that, we need to know how much stoichiometry says she should have gotten:

$$\text{Mass of } HCl = 1.01 \text{ amu} + 35.45 \text{ amu} = 36.46 \text{ amu}$$

$$1 \text{ mole } HCl = 36.46 \text{ g } HCl$$

$$\frac{125 \cancel{\text{ g } HCl}}{1} \times \frac{1 \text{ mole } HCl}{36.46 \cancel{\text{ g } HCl}} = 3.43 \text{ moles } HCl$$

$$\frac{3.43 \cancel{\text{ moles } HCl}}{1} \times \frac{1 \text{ mole } CaCl_2}{2 \cancel{\text{ moles } HCl}} = 1.72 \text{ moles } CaCl_2$$

$$\text{Mass of } CaCl_2 = 40.08 \text{ amu} + 2\times 35.45 \text{ amu} = 110.98 \text{ amu}$$

$$1 \text{ mole } CaCl_2 = 110.98 \text{ g } CaCl_2$$

$$\frac{1.72\ \cancel{\text{moles CaCl}_2}}{1} \times \frac{110.98\ \text{g CaCl}_2}{1\ \cancel{\text{mole CaCl}_2}} = 191\ \text{g CaCl}_2$$

That's what stoichiometry says he should get. However, he only got 85.6% of that.

$$\text{Percent Yield} = (\text{Actual Yield} \div \text{Theoretical Yield}) \times 100\%$$

$$85.6\% = (\text{Actual Yield} \div 191\ \text{g}) \times 100\%$$

Dividing both sides of the equation by 100% and multiplying both sides by 191 g gives us:

$$\frac{85.6\cancel{\%}}{100\cancel{\%}} \times 191\,\text{g} = \text{Actual Yield}$$

$$163\ \text{g} = \text{Actual Yield}$$

3. Empirical formulas have no common factors in their numbers. That means <u>(b) and (d) are empirical formulas</u>.

4. Choice (a) in #3 has a common factor of 2. When we divide by that, we get $\underline{NO_2}$. Choice (c) has a common factor of 2 in its numbers, and when we divide by that, we get $\underline{C_2H_5}$. Choice (e) has a common factor of 3 in its numbers, and when we divide by that, we get $\underline{C_2H_2O}$.

5. The molecular formula is $\underline{C_4H_8}$. It's possible the empirical formula is the molecular formula, so let's see what the mass of the empirical formula is:

$$\text{Mass of } CH_2 = 12.01\ \text{amu} + 2\times1.01\ \text{amu} = 14.03\ \text{amu}$$

This tells us the molar mass is 14.03 g. That's not the molar mass of the substance. However, if we divide the molar mass of the substance by the mass of the empirical formula, we get:

$$56.12 \div 14.03 = 4.000$$

This tells us the molecular formula is four times the empirical formula, or C_4H_8.

6. The empirical formula is $\underline{Fe_2O_3}$. Combustion means chemically reacting with oxygen. If there is only one product, that means the iron and oxygen must both be in that single product. The conservation of mass tells us that all 164.5 grams of product had to come from the reactants, but only 115.0 g of it is iron. The rest, therefore, has to be oxygen:

$$\text{Mass of oxygen} = 164.5\ \text{g} - 115.0\ \text{g} = 49.5\ \text{g}$$

Remember, this is subtraction, so we go by decimal place. Both numbers have their last significant figure in the tenths place, so our answer must be written to the tenths place. Now we can figure out the moles of each. Remember, we are worried about how many moles are in the resulting *compound*, so the fact that oxygen is a homonuclear diatomic is not relevant. We need to know the moles of iron atoms and oxygen atoms in the compound:

$$\frac{115.0 \cancel{\text{ g Fe}}}{1} \times \frac{1 \text{ mole Fe}}{55.85 \cancel{\text{ g Fe}}} = 2.059 \text{ moles Fe}$$

$$\frac{49.5 \cancel{\text{ g O}}}{1} \times \frac{1 \text{ mole O}}{16.00 \cancel{\text{ g O}}} = 3.09 \text{ moles O}$$

The smallest number of moles is for iron, so we can divide by it to get the ratio of O to Fe:

$$\text{Ratio of O atoms to Fe atoms} = \frac{3.09 \cancel{\text{ moles}}}{2.059 \cancel{\text{ moles}}} = 1.50$$

The empirical formula, then, is $FeO_{1.5}$. Of course, we need to get rid of the decimal, so if we multiply both elements by 2, we get an empirical formula of Fe_2O_3.

7. The empirical formula is $\underline{CH_4}$. In a combustion analysis, the moles of carbon dioxide tell us how many moles of carbon are in the original compound:

$$\text{Mass of } CO_2 = 12.01 \text{ amu} + 2 \times 16.00 \text{ amu} = 44.01 \text{ amu}$$

$$1 \text{ mole } CO_2 = 44.01 \text{ g } CO_2$$

$$\frac{412.6 \cancel{\text{ g } CO_2}}{1} \times \frac{1 \text{ mole } CO_2}{44.01 \cancel{\text{ g } CO_2}} = 9.375 \text{ moles } CO_2$$

Since each CO_2 has one C atom, that means there are 9.375 moles of carbon atoms in the original compound.

The moles of water will tell us the moles of hydrogen in the compound:

$$\text{Mass of } H_2O = 2 \times 1.01 \text{ amu} + 16.00 \text{ amu} = 18.02 \text{ amu}$$

$$1 \text{ mole } H_2O = 18.02 \text{ g } H_2O$$

$$\frac{337.9 \cancel{\text{ g } H_2O}}{1} \times \frac{1 \text{ mole } H_2O}{18.02 \cancel{\text{ g } H_2O}} = 18.75 \text{ moles } H_2O$$

For every H_2O, there are two H atoms. That means there are $2 \times 18.75 = 37.50$ moles of H atoms in the compound. The problem says there are only C and H atoms in the compound, so we are done. That means we don't need to use the mass of the liquid, even though it is given. We can now get the ratio. Carbon has the smallest number of moles, so we divide by it:

$$\text{Ratio of H atoms to C atoms} = \frac{37.50\ \cancel{\text{moles}}}{9.375\ \cancel{\text{moles}}} = 4.000$$

That's an integer, so we don't need to round. Thus, we have CH_4.

8. The empirical formula is $\underline{C_2H_6O}$. Since it was burned in excess oxygen, we know the unknown was the limiting reactant. Thus, the moles of carbon dioxide tells us the moles of carbon atoms in the compound:

$$\frac{47.8\ \cancel{\text{g CO}_2}}{1} \times \frac{1\ \text{mole CO}_2}{44.01\ \cancel{\text{g CO}_2}} = 1.09\ \text{moles CO}_2$$

The moles of water tell us the moles of hydrogen atoms in the compound:

$$\frac{29.3\ \cancel{\text{g H}_2\text{O}}}{1} \times \frac{1\ \text{mole H}_2\text{O}}{18.02\ \cancel{\text{g H}_2\text{O}}} = 1.63\ \text{moles H}_2\text{O}$$

Since there are two H atoms for every water molecule, that means there are 2×1.63 = 3.26 moles of H atoms in the substance.

We can't figure out the empirical formula yet, however, because the problem didn't tell us the substance contains only C and H atoms. Carbon dioxide and water were the only products, but they also have oxygen in them, so there could have been oxygen in the compound. To figure that out, we have to see how much mass can be accounted for by the carbon and hydrogen:

$$\frac{1.09\ \cancel{\text{moles C}}}{1} \times \frac{12.01\ \text{g C}}{1\ \cancel{\text{mole C}}} = 13.1\ \text{g C}$$

$$\frac{3.26\ \cancel{\text{moles H}}}{1} \times \frac{1.01\ \text{g H}}{1\ \cancel{\text{mole H}}} = 3.29\ \text{g H}$$

Those add up to 16.4 grams, but the sample had a mass of 25.0 g. That means the remaining mass must have been from oxygen:

$$\text{Mass of oxygen in sample} = 25.0\ \text{g} - 16.4\ \text{g} = 8.6\ \text{g}$$

Now we need to find the moles of oxygen that corresponds to:

$$\frac{8.6\ \cancel{\text{g O}}}{1} \times \frac{1\ \text{mole O}}{16.00\ \cancel{\text{g O}}} = 0.54\ \text{moles O}$$

Oxygen has the smallest number of moles, so it gets a "1," and we divide the others by it:

$$\text{Ratio of C atoms to O atoms} = \frac{1.09\ \cancel{\text{moles}}}{0.54\ \cancel{\text{moles}}} = 2.0$$

$$\text{Ratio of H atoms to O atoms} = \frac{3.26\ \cancel{\text{moles}}}{0.54\ \cancel{\text{moles}}} = 6.0$$

That means the empirical formula is C_2H_6O.

9. The compound is <u>31.91% K, 28.93% Cl, and 39.17% O</u>. To get the percent composition, we have to divide the mass of each element by the total and then multiply by 100%:

$$\text{Mass of } KClO_3 = 39.10 \text{ amu} + 35.45 \text{ amu} + 3\times16.00 \text{ amu} = 122.55 \text{ amu}$$

To get the percent composition for each element, we just divide the mass of each element in the compound by the total mass of the compound and then multiply by 100%.

$$\text{Percent K} = \frac{39.10\ \cancel{\text{amu}}}{122.55\ \cancel{\text{amu}}} \times 100\% = 31.91\%$$

$$\text{Percent Cl} = \frac{35.45\ \cancel{\text{amu}}}{122.55\ \cancel{\text{amu}}} \times 100\% = 28.93\%$$

$$\text{Percent O} = \frac{3\times16.00\ \cancel{\text{amu}}}{122.55\ \cancel{\text{amu}}} \times 100\% = 39.17\%$$

The percentages add to 100.01, which is good. There is always some error in the last significant figure.

10. The empirical formula is <u>$SrCO_3$</u>. When we are given percentages, we can assume any mass of the substance and work out the grams from the percentages. It is easiest to assume 100.0 g, that way we know the substance has 59 g of Sr, 8.0 g of C, and 33 g of O.

$$\frac{59\ \cancel{\text{g Sr}}}{1} \times \frac{1 \text{ mole Sr}}{87.62\ \cancel{\text{g Sr}}} = 0.67 \text{ moles Sr}$$

$$\frac{8.0\ \cancel{\text{g C}}}{1} \times \frac{1 \text{ mole C}}{12.01\ \cancel{\text{g C}}} = 0.67 \text{ moles C}$$

$$\frac{33\ \cancel{\text{g O}}}{1} \times \frac{1 \text{ mole O}}{16.00\ \cancel{\text{g O}}} = 2.1 \text{ moles O}$$

We can divide by either the moles of Sr or C, since they are the smallest.

$$\text{Ratio of C atoms to Sr atoms} = \frac{0.67\ \cancel{\text{moles}}}{0.67\ \cancel{\text{moles}}} = 1.0$$

$$\text{Ratio of O atoms to Sr atoms} = \frac{2.1 \cancel{\text{moles}}}{0.67 \cancel{\text{moles}}} = 3.1$$

The 3.1 can just be rounded to 3.

11. a. You are supposed to recognize the ammonium ion as NH_4^+. Since oxygen is in group 6A, it takes on a 2- charge in ionic compounds. Switching the numbers and dropping the charges gives us $\underline{(NH_4)_2O}$.

b. Sodium is in Group 1A, so it becomes Na^+ in an ionic compound. You are supposed to recognize the sulfate ion as SO_4^{2-}. Switching the numbers and dropping the charges gives us $\underline{Na_2SO_4}$. We don't need parentheses around the sulfate ion, because there is only one.

c. Magnesium is in group 2A, so it is Mg^{2+} in ionic compounds. You are supposed to recognize the phosphate ion as PO_4^{3-}. Switching the numbers and dropping the charges give us $\underline{Mg_3(PO_4)_2}$.

d. The Roman numeral tells us the cobalt ion in this compond is Co^{2+}. You are supposed to recognize the carbonate ion as CO_3^{2-}. The numbers are equal, so we ignore them: $\underline{CoCO_3}$.