Review (AP Chem)
Chapter 3-Mass Relationships in Chemical Reactions

- **In text book**: p111 # 104, 107, 119, 131 *(HINT: $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2\ HCl$)*, 134, 136

104 The carat is a unit of mass used by jewelers. One carat is exactly 200 mg. How many carbon atoms are in a 24-carat diamond?

$$24\text{ carat} \times \frac{200\text{ mg C}}{1\text{ carat}} \times \frac{0.001\text{ g C}}{1\text{ mg C}} \times \frac{1\text{ mol C}}{12.01\text{ g C}} \times \frac{6.022 \times 10^{23}\text{ atoms C}}{1\text{ mol C}} = \mathbf{2.4 \times 10^{23}\text{ atoms C}}$$

107 An impure sample of Zn is treated with an excess of H_2SO_4 to form $ZnSO_4$ and H_2.

(a) Write the balanced equation for the reaction.

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

(b) If 0.0764 g of H_2 is obtained from 3.86 g of Zn, calculate the % purity of the Zn.

We assume that a pure sample would produce the theoretical yield of H_2.

$$3.86\text{ g Zn} \times \frac{1\text{ mol Zn}}{65.39\text{ g Zn}} \times \frac{1\text{ mol H}_2}{1\text{ mol Zn}} \times \frac{2.016\text{ g H}_2}{1\text{ mol H}_2} = 0.119\text{ g H}_2$$

$$\textbf{percent purity} = \frac{0.0764\text{ g H}_2}{0.119\text{ g H}_2} \times 100\% = \mathbf{64.2\%}$$

(c) What assumptions must you make in (b)?

We assume that the impurities are inert and do not react with the sulfuric acid to produce H_2.

131 The formula of a hydrate of barium chloride is $BaCl_2 \cdot xH_2O$. If 1.936 g of the compound gives 1.864 g of $BaSO_4$ upon treatment with sulfuric acid, calculate the value of *x*.

Upon treatment with sulfuric acid, $BaCl_2$ dissolves, losing its waters of hydration.

$$BaCl_2(aq) + H_2SO_4(aq) \longrightarrow BaSO_4(s) + 2HCl(aq)$$

Next, calculate the mass of anhydrous $BaCl_2$ based on the amount of $BaSO_4$ produced.

$$1.864\text{ g BaSO}_4 \times \frac{1\text{ mol BaSO}_4}{233.4\text{ g BaSO}_4} \times \frac{1\text{ mol BaCl}_2}{1\text{ mol BaSO}_4} \times \frac{208.2\text{ g BaCl}_2}{1\text{ mol BaCl}_2} = 1.663\text{ g BaCl}_2$$

Mass of water = (1.936 g − 1.663 g) = 0.273 g H_2O.

$$0.273\text{ g H}_2\text{O} \times \frac{1\text{ mol H}_2\text{O}}{18.02\text{ g H}_2\text{O}} = 0.0151\text{ mol H}_2\text{O}$$

$$1.663\text{ g BaCl}_2 \times \frac{1\text{ mol BaCl}_2}{208.2\text{ g BaCl}_2} = 0.00799\text{ mol BaCl}_2$$

The ratio of the number of moles of H_2O to the number of moles of $BaCl_2$ is 0.0151/0.00799 = **1.89**. We round this number to **2**, which is the value of *x*. The formula of the hydrate is $\mathbf{BaCl_2 \cdot 2H_2O}$.

134 When 0.273 g of Mg is heated strongly in an N_2 atmosphere, a chemical reaction occurs. The product has a mass of 0.378 g. Calculate the empirical formula of the compound, containing only Mg and N and name it.

? g N = 0.378 g Mg_xN_y – 0.273 g Mg = 0.105 g N

$$0.273 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} = 0.0112 \text{ mol Mg} \div 0.00749 = 1.5$$

$$0.105 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.00749 \text{ mol N} \div 0.00749 = 1$$

Multiplying by a factor of 2 gives the empirical formula **Mg_3N_2**. The name of this compound is **magnesium nitride**.

136 The anti-knock additive in leaded gasoline contains only C, H, and Pb. When 51.36 g of this compound is burned in an apparatus like that in Figure 3.6, 55.90 g of CO_2 and 28.61 g of H_2O are produced. Determine the empirical formula of the additive.

Step 1: Calculate the mass of C in 55.90 g CO_2, and the mass of H in 28.61 g H_2O.

$$? \text{ g C} = 55.90 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 15.25 \text{ g C}$$

Similarly,

$$? \text{ g H} = 28.61 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 3.201 \text{ g H}$$

Step 2: Calculate the mass of Pb by difference.

$$\text{mass Pb} = 51.6 \text{ g} - (15.25 \text{ g C} + 3.201 \text{ g H}) = 32.91 \text{ g Pb}$$

Step 3: Calculate the number of moles of each element present in the sample and divide by the smallest amount to give the empirical formula.

$$15.25 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 1.270 \text{ mol C} \div 0.1588 \approx 8$$

$$3.201 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 3.176 \text{ mol H} \div 0.1588 \approx 20$$

$$32.91 \text{ g Pb} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} = 0.1588 \text{ mol Pb} \div 0.1588 = 1$$

This gives the empirical formula, **PbC_8H_{20}**. [This is tetra-ethyl lead, $Pb(C_2H_5)_4$]

- **Additional Questions:**

a) 0.755 g sample of hydrated copper(II) sulfate is heated carefully until it had changed completely to anhydrous copper(II) sulfate with a mass of 0.483 g. Determine the value of x in the formula of the hydrate, $CuSO_4 \cdot x\, H_2O$. *(What would you have done in the lab to be sure that no water was left in the sample after heating?)*

First determine the number of moles of anhydrate:

$$? \text{ moles } CuSO_4 = 0.483 \text{ g } CuSO_4 \times \frac{1 \text{ mol } CuSO_4}{159.62 \text{ g } CuSO_4} = 0.00303 \text{ mol } CuSO_4$$

Next, determine the moles of H_2O driven off by heating:

$$? \text{ g } H_2O = 0.755 \text{ g } CuSO_4 \bullet x\, H_2O - 0.483 \text{ g } CuSO_4 = 0.272 \text{ g } H_2O$$

$$? \text{ mol } H_2O = 0.272 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 0.0151 \text{ mol } H_2O$$

Finally, find the ratio $x = \dfrac{\text{mol } H_2O}{\text{mol } CuSO_4} = \dfrac{0.0151 \text{ mol } H_2O}{0.00303 \text{ mol } CuSO_4} = 4.98 \approx 5$

so the hydrate is $\mathbf{CuSO_4 \cdot 5\, H_2O}$, copper(II) sulfate pentahydrate.

To be sure no water remained after heating, you could heat the sample several times and mass it after each heating, until the mass after heating remained constant, indicating that no more water was present to drive off.

b) Write the balanced chemical equation for the following reactions:

1) calcium metal reacts with water to produce aqueous calcium hydroxide and hydrogen gas.
$Ca(s) + 2\, H_2O\, (\ell) \rightarrow Ca(OH)_2\, (aq) + H_2\, (g)$

2) barium hydroxide reacts with sulfuric acid to produce barium sulfate and water.
$Ba(OH)_2\, (s) + H_2SO_4\, (aq) \rightarrow BaSO_4\, (s) + 2\, H_2O\, (\ell)$

3) iron(III) sulfide reacts with hydrogen chloride to form iron(III) chloride and hydrogen sulfide.
$Fe_2S_3\, (s) + 6\, HCl\, (g) \rightarrow 2\, FeCl_3\, (s) + 3\, H_2S\, (g)$

4) carbon disulfide reacts with ammonia to produce hydrogen sulfide and ammonium thiocyanate.
$CS_2\, (\ell) + 2\, NH_3\, (g) \rightarrow H_2S\, (g) + NH_4SCN\, (s)$

c) Hydrogen peroxide can be produced by the following reaction:

$$BaO_2\, (s) + 2\, HCl\, (aq) \rightarrow H_2O_2\, (aq) + BaCl_2\, (aq)$$

- What is the theoretical yield (in grams) of hydrogen peroxide when 1.50g of barium peroxide is treated with 25.0 mL of a hydrochloric acid solution containing 0.0272g of HCl per mL?
- How many grams of the excess reactant are left unreacted?

First we need to determine which reactant is limiting, so determine the moles of H_2O_2 that can be produced from each reactant:

$$? \text{ mol } H_2O_2 = 1.50 \text{ g } BaO_2 \times \frac{1 \text{ mol } BaO_2}{169.33 \text{ g } BaO_2} \times \frac{1 \text{ mol } H_2O_2}{1 \text{ mol } BaO_2} = 0.00886 \text{ mol } H_2O_2$$

$$\text{mass HCl} = 25.0 \text{ mL HCl solution} \times \frac{0.0272 \text{ g HCl}}{1 \text{ mL HCl solution}} = 0.680 \text{ g HCl}$$

$$? \text{ mol } H_2O_2 = 0.680 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \times \frac{1 \text{ mol } H_2O_2}{2 \text{ mol HCl}} = 0.00933 \text{ mol } H_2O_2$$

BaO_2 is limiting, so $\text{TY } H_2O_2 = 0.00886 \text{ mol } H_2O_2 \times \frac{34.02 \text{ g } H_2O_2}{1 \text{ mol } H_2O_2} = \mathbf{0.301 \text{ g } H_2O_2}$

HCl is excess, so determine amount needed:

$$? \text{ g HCl needed} = 0.00886 \text{ mol } H_2O_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol } H_2O_2} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = 0.646 \text{ g HCl}$$

? g HCl remaining = 0.680 g HCl – 0.646 g HCl = **0.034 g HCl**

d) Methyl isothiocyanate (MITC), an organosulfur compound which contains only C, H, N, and S, is used in agriculture as a soil fumigant, mainly for protection against fungi and nematodes. Find the empirical formula for MITC if combustion analysis of a 0.2415-g sample gives 0.2907 g CO_2, 0.08926 g H_2O, a mixture of nitrogen oxides, and 0.2116 g SO_2.

First, find the masses of each element in the original sample:

$$? \text{ g C} = 0.2907 \text{ g } CO_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g } CO_2} = 0.07933 \text{ g C}$$

$$? \text{ g H} = 0.08926 \text{ g } H_2O \times \frac{(2)(1.008 \text{ g H})}{18.02 \text{ g } H_2O} = 0.009886 \text{ g H}$$

$$? \text{ g S} = 0.2116 \text{ g } SO_2 \times \frac{32.07 \text{ g S}}{64.07 \text{ g } SO_2} = 0.1059 \text{ g S}$$

$$? \text{ g N} = 0.2145 \text{ g MITC} - (0.07933 \text{ g C} + 0.009886 \text{ g H} + 0.1059 \text{ g S}) = 0.0463 \text{ g N}$$

Next, find the numbers of moles of each element in the sample:

$$? \text{ mol C} = 0.07933 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.006605 \text{ mol C}$$

$$? \text{ mol H} = 0.009986 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.009907 \text{ mol H}$$

$$? \text{ mol N} = 0.0463 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.00330 \text{ mol N}$$

$$? \text{ mol S} = 0.1059 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 0.003302 \text{ mol S}$$

Dividing each quantity by 0.00330 mol gives $\mathbf{C_2H_3NS}$ as the empirical formula.