Chemistry 161a – Fall 2018 **Chapter 3 Discussion – Extra Key**

1. Some fundamental calculations:
   a. Calculate the number of oxygen atoms in a 288 amu sample.
   $$288\ amu \times \frac{1\ atom\ \text{O}}{16.00\ amu} = 18\ atoms\ \text{O}$$
   b. Calculate the mass (in amu) of 51 Na atoms.
   $$51\ atoms\ \text{Na} \times \frac{22.99\ amu}{1\ atom\ \text{Na}} = 1173\ amu$$
   c. Calculate the number of moles in a sample with 5.00 × 10^{20} atoms of Cr.
   $$5.00 \times 10^{20}\ atoms\ \text{Cr} \times \frac{1\ mol}{6.022 \times 10^{23}\ atom} = 8.30 \times 10^{-4}\ mol\ \text{Cr}$$
   d. Calculate the mass of the sample in part (c).
   $$5.00 \times 10^{20}\ atoms\ \text{Cr} \times \frac{1\ mol}{6.022 \times 10^{23}\ atom} \times \frac{52.00\ g\ \text{Cr}}{1\ mol\ \text{Cr}} = 4.32 \times 10^{-2}\ g\ \text{Cr}$$
   e. How many atoms of Si are in a 5.68 mg sample?
   $$5.68\ mg\ \text{Si} \times \frac{1\ g}{1000\ mg} \times \frac{1\ mol\ \text{Si}}{28.09\ g\ \text{Si}} \times \frac{6.022 \times 10^{23}\ atom}{1\ mol} = 1.22 \times 10^{-20}\ atoms\ \text{Si}$$
   f. How many grams of aluminum sulfate are in a 0.630 mol sample?
   $$0.630\ mol\ \text{Al}_2(\text{SO}_4)_3 \times \frac{342.14\ g\ \text{Al}_2(\text{SO}_4)_3}{1\ mol\ \text{Al}_2(\text{SO}_4)_3} = 216\ g\ \text{Al}_2(\text{SO}_4)_3$$
   g. How many moles are in 50.0 g sample of ammonium carbonate?
   $$50.0\ g\ (\text{NH}_4)_2\text{CO}_3 \times \frac{1\ mol\ (\text{NH}_4)_2\text{CO}_3}{96.094\ g\ (\text{NH}_4)_2\text{CO}_3} = 0.520\ g\ (\text{NH}_4)_2\text{CO}_3$$
   h. What is the mass of one molecule of dinitrogen tetroxide?
   $$1\ molecule\ \text{N}_2\text{O}_4 \times \frac{1\ mol}{6.022 \times 10^{23}\ molecules} \times \frac{92.02\ g\ \text{N}_2\text{O}_4}{1\ mol\ \text{N}_2\text{O}_4} = 1.528 \times 10^{-22}\ g\ \text{N}_2\text{O}_4$$

2. Translate the following descriptions into a <u>balanced</u> chemical equation.
   a. Solid iron metal reacts with oxygen gas to produce solid iron (III) oxide.
   $$4\text{Fe}\ (s) + 3\text{O}_2\ (g) \rightarrow 2\text{Fe}_2\text{O}_3\ (s)$$
   b. The combustion of solid iron (II) oxide produces solid iron (III) oxide.
   $$4\text{FeO}\ (s) + 1\text{O}_2\ (g) \rightarrow 2\text{Fe}_2\text{O}_3\ (s)$$
   c. Solid potassium metal reacts with water to make hydrogen gas and aqueous potassium hydroxide.
   $$2\text{K}\ (s) + 2\text{H}_2\text{O}\ (l) \rightarrow 1\text{H}_2\ (g) + 2\text{KOH}\ (aq)$$
   d. Propane gas combusts to give off carbon dioxide and water vapor.
   $$1\text{C}_3\text{H}_8\ (g) + 5\text{O}_2\ (g) \rightarrow 3\text{CO}_2\ (g) + 4\text{H}_2\text{O}\ (g)$$
   e. Dihydrogen sulfide gas is bubbled through an aqueous solution of lead (II) nitrate and solid lead (II) sulfide forms alongside aqueous nitric acid.
   $$1\text{H}_2\text{S}\ (g) + 1\text{Pb(NO}_3)_2\ (aq) \rightarrow 1\text{PbS}\ (s) + 2\text{HNO}_3\ (aq)$$

f. Sulfuric acid is poured onto solid aluminum to give off hydrogen gas and a solution of aluminum sulfate.

$$3H_2SO_4\ (aq) + 2Al\ (s) \rightarrow 3H_2\ (g) + 1Al_2(SO_4)_3\ (aq)$$

g. A copper wire dipped in a solution of silver (I) nitrate produces silver metal and copper (II) nitrate solution.

$$1Cu\ (s) + 2AgNO_3\ (aq) \rightarrow 2Ag\ (s) + 1Cu(NO_3)_2\ (aq)$$

3. If you have equal masses of each compound (H_2SO_4, $C_{12}H_{22}O_{11}$, $KClO_3$), which sample has the greatest number of oxygen atoms?
(Note: There are two ways you can do this problem!)
Method 1: Determine mass percent of O in each. → *H_2SO_4: 65 % O | $C_{12}H_{22}O_{11}$: 51% O | $KClO_3$: 39% O*
Method 2 (much longer): Determine number of O atoms in each compound.

4. Striking a match results in the following <u>unbalanced</u> chemical reaction:

$$\_\_\_\_\ P_4\ (s) + \_\_\_\_\ KClO_3\ (s) \rightarrow \_\_\_\_\ KCl\ (s) + \_\_\_\_\ P_2O_5\ (s)$$

a. Balance the equation above.

$$3P_4\ (s) + 10KClO_3\ (s) \rightarrow 10KCl\ (s) + 6P_2O_5\ (s)$$

b. If 15.0 mg of P_2O_5 was produced in this reaction, what masses of P_4 and $KClO_3$ were required?
Start by converting 15.0 mg of P_2O_5 to mol of P_2O_5:

$$15.0\ mg\ \mathrm{P_2O_5} \times \frac{1\ mg}{1000\ mg} \times \frac{1\ mol\ \mathrm{P_2O_5}}{141.94\ g\ \mathrm{P_2O_5}} = 1.057 \times 10^{-4}\ mol\ \mathrm{P_2O_5}$$

Now determine how much P_4 is needed to produce this much P_2O_5:

$$1.057 \times 10^{-4}\ mol\ \mathrm{P_2O_5} \times \frac{3\ mol\ \mathrm{P_4}}{6\ mol\ \mathrm{P_2O_5}} \times \frac{123.88\ g\ \mathrm{P_4}}{1\ mol\ \mathrm{P_4}} = 6.55 \times 10^{-3}\ g\ \mathrm{P_4}$$

Now determine how much $KClO_3$ is needed to produce this much P_2O_5:

$$1.057 \times 10^{-4}\ mol\ \mathrm{P_2O_5} \times \frac{10\ mol\ \mathrm{KClO_3}}{6\ mol\ \mathrm{P_2O_5}} \times \frac{122.55\ g\ \mathrm{KClO_3}}{1\ mol\ \mathrm{KClO_3}} = 2.16 \times 10^{-2}\ g\ \mathrm{KClO_3}$$

5. You react 10.0 g of hydrogen gas with 60.0 g of oxygen gas to form water. Determine the amount of water formed and the amount of excess reactant (both in grams) after the reaction is complete.

Balanced chemical equation: $2H_2\ (g) + 1O_2\ (g) \rightarrow 1H_2O\ (g)$

Now determine whether the limiting reactant is H_2 or O_2 gas (two ways):

Method 1:	Method 2:
How much H_2O can we make from all H_2? $10.0\ g\ \mathrm{H_2} \times \frac{1\ mol\ \mathrm{H_2}}{2.016\ g\ \mathrm{H_2}} \times \frac{2\ mol\ \mathrm{H_2O}}{2\ mol\ \mathrm{H_2}} = 4.96\ mol\ \mathrm{H_2O}$ *How much H_2O can we make from all O_2?* $60.0\ g\ \mathrm{O_2} \times \frac{1\ mol\ \mathrm{O_2}}{32.00\ g\ \mathrm{O_2}} \times \frac{2\ mol\ \mathrm{H_2O}}{1\ mol\ \mathrm{O_2}} = 3.75\ mol\ \mathrm{H_2O}$ *Therefore, O_2 is limiting (makes less H_2O)!*	*We have 4.96 mol H_2 and 1.875 mol O_2. How much H_2 do we need to react with all of the O_2?* $60.0\ g\ \mathrm{O_2} \times \frac{1\ mol\ \mathrm{O_2}}{32.00\ g\ \mathrm{O_2}} \times \frac{2\ mol\ \mathrm{H_2}}{1\ mol\ \mathrm{O_2}} = 3.75\ mol\ \mathrm{H_2}$ *We <u>need</u> 3.75 mol H_2, but <u>have</u> 4.96 mol H_2, so H_2 is in excess.* *Therefore, O_2 is limiting!*

H_2 is in excess and we need to use 3.75 mol. Leftover = 4.96 mol – 3.75 mol = 1.21 mol H_2 = 2.44 g H_2.
From method 1, we know we make 3.75 mol H_2O or 67.6 g H_2O.

6. Consider mixing an excess of lead (II) nitrate (*aq*) with 0.0800 mol sodium chloride. Determine the mass of solid lead (II) chloride formed assuming a complete reaction.

   Balanced chemical equation: $1Pb(NO_3)_2\ (aq) + 2NaCl\ (aq) \rightarrow 1PbCl_2\ (s) + 2NaNO_3\ (aq)$

   Since we are told $Pb(NO_3)_2$ is in excess, the NaCl is the limiting reactant.

$$0.0800\ mol\ \mathrm{NaCl} \times \frac{1\ mol\ \mathrm{PbCl_2}}{2\ mol\ \mathrm{NaCl}} \times \frac{278.10\ g\ \mathrm{PbCl_2}}{1\ mol\ \mathrm{PbCl_2}} = 11.1\ g\ \mathrm{PbCl_2}$$

7. The percent by mass of nitrogen is 46.7% for a species containing only nitrogen and oxygen. Which of the following could this species be?

   N_2O_5 N_2O NO_2 NO NO_3

   Assuming a 100 g sample of N_xO_y compound: 46.7 g N and 53.3 g O
   Determine the empirical formula:

$$46.7\ g\ \mathrm{N} \times \frac{1\ mol\ \mathrm{N}}{14.01\ g\ \mathrm{N}} = 3.333\ mol\ \mathrm{N} \quad \Longrightarrow \quad \frac{3.333\ mol\ \mathrm{N}}{3.331} = 1\ mol\ \mathrm{N}$$

$$53.3\ g\ \mathrm{O} \times \frac{1\ mol\ \mathrm{O}}{16.00\ g\ \mathrm{O}} = 3.331\ mol\ \mathrm{O} \quad \Longrightarrow \quad \frac{3.331\ mol\ \mathrm{O}}{3.331} = 1\ mol\ \mathrm{O}$$

   Empirical formula: NO

8. If 5.00 g of CH_4 is burned, what mass of water can be produced?

   Balanced chemical equation: $1CH_4\ (g) + 2O_2\ (g) \rightarrow 1CO_2\ (g) + 2H_2O\ (g)$

$$5.00\ g\ \mathrm{CH_4} \times \frac{1\ mol\ \mathrm{CH_4}}{16.04\ g\ \mathrm{CH_4}} \times \frac{2\ mol\ \mathrm{H_2O}}{1\ mol\ \mathrm{CH_4}} \times \frac{18.02\ g\ \mathrm{H_2O}}{1\ mol\ \mathrm{H_2O}} = 11.2\ g\ \mathrm{H_2O}$$