Name: ______________________________________________ Date: _____________________ CHEM 1A

*Monday/Tuesday, October 14 & 15, 2019 - Stoichiometry (Chapter 3) Part 2*

**I. Warm-Up**

1. How many significant figures does 0.000123045560 have? 9
2. 12.5849 / 2.4 = 5.2
3. 432.5-24.3984=408.1
4. 12.0(11.90-11.8) =133.
5. $\frac{1.203\times10^{6}}{0.000360-2.40\times10^{5}}$ = -5.01

**Significant Figures Rules**

Addition/Subtraction –The answer is precise as the least accurate number

Multiplication/Division – The answer has the same sig figs as the number with the least sig. figs.

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**II. Limiting Reagents and Theoretical Yields**

1. Consider the following unbalanced reaction – assume the reaction goes to completion in each scenario:

   $NH_3 + O_2 \rightarrow NO_2 + H_2O$

   a. How many moles of oxygen gas are required to make 12.8 moles of nitrogen dioxide? 22.4 mol $O_2$

   1st balance the reaction ⇒ $4\ NH_3 + 7\ O_2 \rightarrow 4\ NO_2 + 6\ H_2O$

   12.8 mol $NO_2$ x $\frac{7 \text{ mol } O_2}{4 \text{ molN } O_2}$ = 22.4 mol $O_2$

   b. How many grams of water can be produced from 9.64 g of ammonia?

9.64 $NH_3$ x $\frac{1 \text{ mol } NH_3}{17.034 \text{ g}}$ x $\frac{6 \text{ mol } H_2O}{4 \text{ mol } NH_3}$ x $\frac{18.016 \text{ g} H_2O}{1 \text{ mol } H_2O}$ = 15.3 g $H_2O$

   c. Identify the limiting reagent if 3 moles of ammonia is combined with 5 moles of oxygen. $O_2$ is LR

   3 $\frac{\text{molN } H_3}{4}$ = 0.75 vs. $\frac{5 \text{ mol } O_2}{7}$ = 0.71 ⇒ smaller so $O_2$ is the LR

   d. Identify the limiting reagent if 10.00 g of ammonia is combined with 28.00 g of oxygen. $O_2$ is the LR

   10 g $NH_3$ x $\frac{1 \text{ mol } NH_3}{17.034 \text{ g}}$ = 0.587 mol $NH_3$ ⇒ $\frac{0.587 \text{ mol } NH_3}{4}$ = 0.15 vs.

   28 g $O_2$ x $\frac{1 \text{ mol } O_2}{16 \text{ g}}$ = 0.875 mol $O_2$ ⇒ $\frac{0.875 \text{ mol } O_2}{7}$ = 0.12 ⇒ smaller so $O_2$ is the LR

   e. How many grams of each species will be present if 10.00 g of ammonia is combined with 28.00 g oxygen?

Since $O_2$ is the LR there will be none of it left and $O_2$ will determine the answer for all the remaining species

$NH_3$ ⇒ 0.875mol $O_2$ x $\frac{4 \text{ mol } NH_3}{7 \text{ mol } O_2}$ x $\frac{17.034 \text{ g}}{1 \text{ mol } NH_3}$ = 8.5 g of $NH_3$ is consumed ⇒ 10 g – 8.5 g = 1.5 g $NH_3$ remains

$NO_2$ ⇒ 0.875mol $O_2$ x $\frac{4 \text{ mol } NO_2}{7 \text{ mol } O_2}$ x $\frac{46.01 \text{ g}}{1 \text{ mol } NO_2}$ =23 g of $NO_2$ is made

$H_2O$ ⇒ 0.875mol $O_2$ x $\frac{6 \text{ mol } H_2O}{7 \text{ mol } O_2}$ x $\frac{18.016 \text{ g}}{1 \text{ mol } H_2O}$ = 13.5 g of $H_2O$ is made

Also try doing the ICF Table method.

2. Consider the following unbalanced reaction:

$$P_4O_{10}\ (s) + PCl_5\ (g) \longrightarrow POCl_3\ (g)$$

When 35 g of solid $P_4O_{10}$ and 42 g of gaseous $PCl_5$ is combined 47 g of $POCl_3$ is produced. What is the percent yield for this process?

Balanced rxn⇒ $P_4O_{10}$ (s) + 6 $PCl_5$ (g) → 10 $POCl_3$ (g)

ID the LR ⇒ 35 g $P_4O_{10}$ x $\frac{1\text{ mole}}{283.88\text{ g}}$ = 0.1233 mol ⇒$\frac{0.1233\text{ mol}}{1}$ = 0.1233 vs.

42 g $PCl_5$x $\frac{1\text{ mole}}{208.22\text{ g}}$ = 0.202 mol ⇒$\frac{0.202\text{ mol}}{6}$ = 0.0337 ⇒ smaller so $PCl_5$ is the LR

Theoretical yield ⇒ 0.202 mol $PCl_5$ x $\frac{10\text{ mol }POCl_3}{6\text{ mol }PCl_5}$ x $\frac{153.32\text{ g }POCl_3}{1\text{ mol }POCl_3}$ = 51.62 g $POCl_3$

Percent Yield = $\frac{47\text{ g}}{51.62\text{ g}}$ x 100 = 91 %

III. **Composition of Compounds**

*Recall: Law of Def. & Multiple Prop.*

**EmpiricalFormula**–Tells the ratio of atoms of each element present in the compound.

e.g. –$CH_2O$ is the empirical formula

1. *Convert given values into moles for each element*
2. *Divide all moles by the smallest mole value*
3. *If you have all whole numbers you have the EF – if not try multiplying them all by 2 or 3 etc.*

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**Molecular Formula** – Tells the actual numbers of atoms of each element in a molecule.

e.g. – $C_6H_{12}O_6$ is the molecular formula of glucose

1. *Derive empirical formula*
2. *Determine the empirical mass*
3. $\frac{(Molar\ \ mass)}{(empirical\ \ mass)}$ *= multiple*
4. *Multiply the empirical formula by the multiple*

3. Calculate the mass percent of Cl in NaCl.

35.5g Cl/(35.5g + 23g) NaCl = 61% Cl

4. A 1.40 g sample of silicon reacts with fluorine to produce 5.2 g of a product. What is the empirical formula of the compound?

1.40g Si + x g $F_2$ →5.2g (???)

By conservation of mass, 5.2g – 1.4g = x = 3.8g

Convert the grams to moles of substance:

$$1.4g\ Si \times \frac{1\ mol\ Si}{28\ g\ Si} = 0.05 mol\ Si;\ \ 3.8g\ F_2 \times \frac{2\ mol\ F}{38\ g\ F_2} = 0.2 mol\ F$$

Divide for ratio 0.05 => 1 mol Si: 4 mol F => EF = $SiF_4$

5. The empirical formula for xylene is $C_4H_5$ and xylene has a molar mass of 106.16 g/mol. Determine the molecular formula for xylene.

Molar Mass(given)/Molar Mass(EF) = 106.16 g/mol/53.08 g/mol= 2. Double the empirical formula =>$C_8H_{10}$

6. An alkali metal oxide contains 83.01% metal by mass. Determine the identity of the metal.

Alkali metal is group 1 ⇒ so Ca and Zn are not possible ⇒ Group 1 are 1+ in ionic compounds and oxides are 2– ⇒ an alkali metal oxide will have the formula ⇒ $M_2O$

83.01 % = $\frac{2\text{ M}}{2\text{M}+16}$ x 100⇒ M is 39.1 g/mol⇒ K

7. Tryptophan is an amino acid that is 64.7% carbon, 5.9% hydrogen, 13.7% nitrogen and 15.7% oxygen. What is the empirical formula for tryptophan?

If you have 100 g sample of tryptophan ⇒ 64.7 g C : 5.9 g H : 13.7 g N : 15.7 g O

$$\frac{64.7\text{ g C}}{12.01\frac{\text{g}}{\text{mol}}} : \frac{5.9\text{ g H}}{1.008\frac{\text{g}}{\text{mol}}} : \frac{13.7\text{ g N}}{14.01\frac{\text{g}}{\text{mol}}} : \frac{15.7\text{ g O}}{16\frac{\text{g}}{\text{mol}}}$$

5.387 molC : 5.853 mol H : 0.9779 mol N : 0.9812 mol O

divide each by the smallest ratio (0.9779)

5.5 molC : 6 mol H : 1 mol N : 1 mol O

double each to get the Empirical Formula

$C_{11}H_{12}N_2O_2$

8. **The Combustion Problem -** A 0.4647-g sample of a compound known to contain only carbon, hydrogen, and oxygen was burned in oxygen to yield 0.8635 g of $CO_2$ and 0.1767 g of $H_2O$. If the molar mass is 213 g/mol, what is the molecular formula of the compound?

$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$ ⇒ all of the C ends up in the $CO_2$ and all of the H ends up in the $H_2O$

the O is not so obvious

C ⇒ $\frac{12.01\text{ g C}}{44.01\text{ g CO}_2} = \frac{\text{x g C}}{0.8635\text{ g CO}_2}$ ⇒ 0.2356 g C x $\frac{1\text{ mol C}}{12.01\text{ g C}}$ = 0.01962 mol C

H ⇒ $\frac{2.016\text{ g C}}{18.016\text{ g H}_2\text{O}} = \frac{\text{x g H}}{0.1767\text{ g H}_2\text{O}}$ ⇒ 0.01977 g H x $\frac{1\text{ mol H}}{1.008\text{ g H}}$ = 0.01961 mol H

O ⇒ 0.4647 g cmpd– 0.2356 g C – 0.01977 g H = 0.2093 g O x $\frac{1\text{ mol O}}{16\text{ g O}}$ = 0.01308 mol O

divide each by the smallest (0.01308)

1.5 molC : 1.5 mol H : 1 mol O

double each ⇒$C_3H_3O_2$ => MW = 71 g/mol

(213 g/mol)/(71 g/mol) = 3 => MF $C_9H_9O_6$

9. **Hmm…** Two compounds contain the same metal and oxygen. Compound I has 13.38% oxygen and Compound 2 has 9.334% oxygen.

a. Calculate the mass of oxygen per 1.000 g of metal.

Assuming 1.00g of metal, Compound I - 0.1338g of oxygen; Compound II – 0.09334g of oxygen

b. If the first compound is $MO_2$ what is the formula of the second compound?

By law of MP, the second compound has 2/3 as much oxygen, so the formula is MO4/3 = M3O4

c. Name the metal.

Looking at the mass of metal for 1 mol, the metal is most likely lead (Pb).