**CHEMISTRY 109 – Help Sheet #3**

**REVIEW (Part III): Stoichiometry (Part I)**

**Review the appropriate topics for your lecture**

Prepared by Dr. Tony Jacob

https://clc.chem.wisc.edu (Resources page)

Nuggets: Mole Calculations; Molar Mass; Mass %; Empirical/Molecular formulas; Finding an unknown metal M

**MOLE** (abbreviated mol): a number equal to 6.022 x 10^23 units = Avogadro’s number; average atomic mass on Periodic Table (number below symbol) equals weight in grams for 1 mol of that element (e.g., 1 mol C = 12.01 g C)

**MOLEcular MASS** = mass of 1 mol of substance in grams

**Example 1:** What is the molar mass of (NH₄)₂O? (AAM = average atomic mass)

**Answer 1:** 52.10 g/mol \( \{2(\text{AAM}_\text{N}) + 8(\text{AAM}_\text{H}) + 1(\text{AAM}_\text{O}) = 2(14.01) + 8(1.01) + 1(16.00) = 52.10 \text{g (NH}_4\text{)}_2\text{O/1mol (NH}_4\text{)}_2\text{O (or just 52.10 g/mol)} \}

**Example 2:** What is the molar mass of CuSO₄·5H₂O? (This is a hydrated compound.)

**Answer 2:** 249.72 g/mol \( \{1(\text{AAM}_\text{Cu}) + 1(\text{AAM}_\text{O}) + 4(\text{AAM}_\text{O}) + 10(\text{AAM}_\text{H}) + 5(\text{AAM}_\text{O}) = 1(63.55) + 1(32.07) + 4(16.00) + 10(1.01) + 5(16.00) = 249.72 \text{g CuSO}_4\cdot5\text{H}_2\text{O/1mol CuSO}_4\cdot5\text{H}_2\text{O (or just 249.72 g/mol)} \} \)

**MASS PERCENT of elements within a compound**

\[
\text{mass}\%_A = \frac{\text{mass}_A}{\text{total mass}} \times 100\%
\]

**Example 3:** What is the mass% of N in Ca(NO₃)₂?

**Answer 3:** 17.1%

1. Assume 1 mol \( \rightarrow \) total mass = molar mass = 1(Ca) + 2(N) + 6(O) = 1(40.1) + 2(14.0) + 6(16.0) = 164.1 g Ca(NO₃)₂
2. Find mass of N in 1 mol of Ca(NO₃)₂: 2 mol N x (14.0 g N/1 mol N) = 28.0 g N
3. mass\%N = \frac{\text{mass N}}{\text{total mass}} \times 100\%; \text{mass}\%N = \frac{28.0}{164.1} \times 100\% = 17.1% (partly)

**EMPIRICAL FORMULA:** the simplest formula that shows the ratio between atoms within the compound

**Determining empirical formula from mass percents:**

1. Assume 100g
2. \% \( \rightarrow \) g
3. g \( \rightarrow \) mol (if g originally given and not mass\% \( \rightarrow \) start at step 3!)
4. Write chemical formula; divide by smallest number of moles
5. Fractions: \( \frac{1}{2} \) (0.5) \( \rightarrow \) x2; \( \frac{1}{3} \) or \( \frac{2}{3} \) (0.33, 0.66) \( \rightarrow \) x3; \( \frac{1}{4} \) or \( \frac{3}{4} \) (0.25, 0.75) \( \rightarrow \) x4

**Example 4:** What is the empirical formula for a compound with 43.64% phosphorous and 56.36% oxygen?

**Answer 4:** P₂O₅

1. assume 100g
2. \% \( \rightarrow \) grams: P: 43.64% x 100g = 43.64 g P; O: 56.36% x 100g = 56.36 g O
3. g \( \rightarrow \) mol: 43.64 g P \( \left( \frac{1 \text{ mol P}}{30.97 g P} \right) = 1.409 \text{ mol P} \); 56.36 g O \( \left( \frac{1 \text{ mol O}}{16.00 g O} \right) = 3.523 \text{ mol O} \)
4. Write EF: 1.409O₃.523 and divide by smallest number of mol: 1.490: \( \frac{1.409}{1.409} = \frac{3.523}{1.409} = \frac{P₂O₅}{P₂O₅} \)
5. Fractions, multiply by 2: \( P \left( \frac{1 \times 2 \text{O}}{2 \times 2.500} \times 2 \right) \rightarrow P₂O₅ = \text{EF}

**MOLECULAR FORMULA:** the exact formula of a compound

For N₂O₄, the molecular formula is N₂O₄ and the empirical formula is NO₂

You can’t determine the molecular formula from mass percent, only the empirical formula

**Determine Molecular Formula from Empirical Formula and the Molar Mass**

1. Determine ratio: molar mass of molecular formula
   molar mass of empirical formula
2. Take empirical formula and multiply each subscript by the number from above ratio
Example 5: Using the prior example (Example 4), what is the molecular formula if the molar mass is 283.88 g/mol?
Answer 5: $P_4O_{10}$

1. Determine ratio:
   \[
   \frac{\text{molar mass}_\text{molecular formula}}{\text{molar mass}_\text{empirical formula}}; \text{ molar mass molecular formula is given in the problem} = 283.88 \text{ g/mol};
   \]

   \[
   \frac{\text{molar mass}_\text{empirical formula}}{\text{molar mass}_\text{molecular formula}} = \frac{283.88 \text{ g/mol}}{141.8 \text{ g/mol}} = 2.01
   \]

2. Multiply empirical formula by ratio: $P(2 \times 2)O(5 \times 2) = P_4O_{10} = \text{molecular formula}$

CONVERTING BETWEEN MOL, GRAMS, AND ATOMS/MOLECULES

Use the flow chart to guide you for basic mole calculations

Example 6: How many grams of $N$ are in 88.0 g $N_2O$?
Answer 6: 56.0 g $N$

\{(This is a grams A \rightarrow grams B calculation; requires 3 conversions (counting the steps in the flow chart); in the calculation below, the fractions in the parentheses have units on the bottom to cancel the prior units and units on the top that you want to end up with:}\n
\[
\frac{88.0 \text{ g} N_2O}{1 \text{ mol} N_2O} \left( \frac{1 \text{ mol} N_2O}{44.0 \text{ g} N_2O} \right) \left( \frac{2 \text{ mol} N}{1 \text{ mol} N_2O} \right) \left( \frac{14.0 \text{ g} N}{1 \text{ mol} N} \right) = 56.0 \text{ g} N
\]

(note: after 1st fraction/parentheses the units are mol A; after 2nd fraction/parentheses the units are g B; after 3rd fraction/parentheses the units are mol B; it is often helpful to cancel out units (not numbers) to check to make sure the set-up is correct: after $\frac{14.0 \text{ g} N}{1 \text{ mol} N}$ cancels $(44.0 \text{ g} N_2O)(2 \text{ mol} N)$)

Example 7: A single molecule weighs $4.98 \times 10^{-23}$ g. What molecule is it?

a. NO$_2$    b. CO$_2$  c. H$_2$O  d. NO    e. CO

Answer 7: d

\{(Determine the molar mass of the molecule and compare to the molar masses of the possible choices.}\n
Step 1. \[
\left( \frac{4.98 \times 10^{-23} \text{ g}}{\text{molecule}} \right) \left( \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} \right) = 29.99 \text{ g/mol}
\]

Step 2. Find molar masses of chemicals given: NO$_2$ = 46 g/mol; CO$_2$ = 44 g/mol; H$_2$O = 18 g/mol; NO = 30 g/mol; CO = 28 g/mol

Unknown molecule is NO which has a molar mass that matches the calculated molar mass from Step 1)

FINDING “M” IN A FORMULA: This is a specific type of mole calculation.

Example 8: 25.00 g of MCl$_3$ is broken into its elements and produces 17.57 g Cl$_2$. What \textit{element} is M?
Answer 8: Sc

\{(Find the molar mass of the unknown metal M using molar mass \(M = \frac{\text{grams of M}}{\text{mol M}}\) and then look it up on the Periodic Table.\}

Step 1. Find grams of M: $\text{g}_{\text{total}} = \text{g}_M + \text{g}_{\text{Cl}} \rightarrow \text{g}_M = 25.00 - 17.57 = 7.43$ g M;

Step 2. Find mol M using g$_{\text{Cl}_2}$ and chemical formula: 17.57 g Cl$_2$ \[
\left( \frac{1 \text{ mol Cl}_2}{70.90 \text{ g Cl}_2} \right) \left( \frac{2 \text{ mol Cl}}{1 \text{ mol Cl}_2} \right) \left( \frac{1 \text{ mol M}}{3 \text{ mol Cl}} \right) = 0.1652 \text{ mol M}
\]

Step 3. molar mass $M = \frac{\text{grams of M}}{\text{mol M}} = \frac{7.43 \text{ g M}}{0.1652 \text{ mol M}} = 44.97$ g/mol \(\rightarrow\) Sc from the periodic table}
Example 9: 1.500g of $M_2CO_3$ was heated releasing 0.623g CO$_2$. What is $M$? (Chemicals with carbonate, CO$_3^{2-}$ are more challenging!)

$$M_2CO_3(s) \rightarrow M_2O(s) + CO_2(g)$$

Answer 9: Na  
\{determine molar mass of $M = g/M$ mol $M$; \\
\text{determine mol } M \text{ using } gCO_2 \text{ and chemical formula: } 0.623g CO_2 \\left( \frac{1mol CO_2}{44.01g CO_2} \right) \\left( \frac{1mol CO_3}{1mol CO_2} \right) \\left( \frac{2mol M}{1mol CO_3} \right) = 0.02831mol M \\
\text{determine } gM \text{ (start by finding gram } CO_3): 0.623g CO_2 \\left( \frac{1mol CO_2}{44.01g CO_2} \right) \\left( \frac{1mol CO_3}{1mol CO_2} \right) \\left( \frac{60.01g CO_3}{1mol CO_3} \right) = 0.8495g CO_3 \\
1.500g = gM + gCO_3 = gM + 0.8495; gM = 0.6505g M; \text{ molar mass } M = \frac{grams M}{mol M} = \frac{0.6505g M}{0.02831mol M} = 22.98g/mol → Na \text{ from the periodic table}\}

1. \text{(Hint: No calculation is needed.)}  
\text{a. Given } 5.0g \text{ samples of } H_2O, N_2O, \text{ and } F_2O, \text{ which of these samples has the most moles?}  
\text{b. Which of these samples has the most molecules?}  
\text{c. Which of these samples has the most atoms?}

2. \text{How many molecules of water will be found in } 5.00\text{ml } H_2O(l)? \text{ Assume the density of water is } 0.997\text{ g/ml.}

3. \text{a. Mercury, Hg, has a molar mass of } 201\text{g/mol. What does one atom of Hg weigh?}  
\text{b. Acetone, } CH_3COCH_3, \text{ has a molar mass of } 58.1\text{g/mol. What does one molecule of acetone weigh?}

4. \text{The mass of a single isotopic version of } O_2 \text{ is measured and found to weigh } 5.978 \times 10^{-23} \text{ grams. Which is} 
\text{the correct isotopic version of } O_2?  
\text{a. } ^{16}O^{16}O \quad \text{b. } ^{16}O^{17}O \quad \text{c. } ^{16}O^{18}O \quad \text{d. } ^{17}O^{18}O \quad \text{e. } ^{18}O^{18}O

5. \text{What is the mass percent of nitrogen in ammonium carbonate, } (NH_4)_2CO_3? \text{?}

6. \text{Which compound has the highest mass percent of oxygen?}  
\text{a. CO} \quad \text{b. CO}_2 \quad \text{c. CO}_3^{2-} \quad \text{d. } H_2O \quad \text{e. all have the same}

7. \text{The composition of adipic acid is } 49.3\% \text{ C, } 6.90\% \text{ H, and } 43.8\% \text{ O and the molar mass is } 146\text{ g/mol. What is} 
\text{the empirical and molecular formula of adipic acid?}

8. \text{15.00g of a metal chloride, } MCl_2, \text{ is completely broken into its elements and produces } 9.583g Cl_2. \text{ Identify} 
\text{the metal, } M.

9. \text{A metal oxide, } M_2O, \text{ is } 93.10\% \text{ by mass metal, } M. \text{ Calculate the atomic mass of the metal.}

\text{ANSWERS}  
1. \text{a. } H_2O \quad \{\text{since } H_2O \text{ has the smallest molar mass and the masses are the same, it will have the most moles}\}  
\text{b. } H_2O \quad \{\text{since } H_2O \text{ has the most moles, it will have the most molecules}\}  
\text{c. } H_2O \quad \{\text{since } H_2O \text{ has the most moles and all the molecules have the same number of atoms per molecule, } H_2O \text{ will have the} 
\text{most atoms}\}

2. \text{1.67 x } 10^{23} \text{ molecules of water} \quad \{D = m/V; m = D(V) = (0.997g/ml)(5.00) = 4.985gH}_2O; 
\text{4.985g } H_2O \left( \frac{1mol H_2O}{18.02g H_2O} \right) \left( \frac{6.022 \times 10^{23} \text{ molecules } H_2O}{1mol H_2O} \right) = 1.666 \times 10^{23} \text{molecules } H_2O \}
3. a. \(3.34 \times 10^{-22} \text{g/atom Hg} \) \(\{\left(\frac{201 \text{g Hg}}{1 \text{mol Hg}}\right)\left(\frac{1 \text{mol Hg}}{6.022 \times 10^{23} \text{atoms Hg}}\right) = 3.338 \times 10^{-22} \text{g/atom Hg}\}\)

more simply: \(\left(\frac{\text{molar mass}}{6.022 \times 10^{23}}\right) = \left(\frac{201 \text{g Hg}}{6.022 \times 10^{23} \text{atoms Hg}}\right) = 3.338 \times 10^{-22} \text{g/atom Hg}\}\)

b. \(9.65 \times 10^{-23} \text{g/molecule CH}_3\text{COCH}_3\)

\(\{\left(\frac{58.1 \text{g CH}_3\text{COCH}_3}{1 \text{mol CH}_3\text{COCH}_3}\right)\left(\frac{1 \text{mol CH}_3\text{COCH}_3}{6.022 \times 10^{23} \text{molecules CH}_3\text{COCH}_3}\right) = 9.648 \times 10^{-23} \text{g/molecule CH}_3\text{COCH}_3;\)

more simply: \(\left(\frac{\text{molar mass}}{6.022 \times 10^{23}}\right) = \left(\frac{58.1 \text{g CH}_3\text{COCH}_3}{6.022 \times 10^{23} \text{molecules CH}_3\text{COCH}_3}\right) = 9.648 \times 10^{-23} \text{g/molecules CH}_3\text{COCH}_3\}\)

4. e \(\{\left(\frac{5.978 \times 10^{-23} \text{g}}{\text{molecule O}_2}\right)\left(\frac{6.022 \times 10^{23} \text{molecules O}_2}{1 \text{mol O}_2}\right) = 36.00 \text{g mole}^{-1} \text{O}_2; \text{now compare the 36 to the mass or mass numbers of the different isotopic versions;} \quad \text{“a” = 32g/mol; “b” = 33g/mol; “c” = 34g/mol; “d” = 35g/mol; “e” = 36g/mol}\}\)

5. \(29.2\% \quad \{\text{mass}\% \text{ N} = (\text{mass N/total mass}) \times 100\%; \text{assume 1 mol} \rightarrow \text{molar mass} = \text{total mass} = 96.09\text{g};\)

\[
\begin{align*}
\text{mass N} &= 1\text{mol (NH}_4)_2\text{CO}_3 \left(\frac{2\text{mol N}}{1\text{mol (NH}_4)_2\text{CO}_3}\right) \left(\frac{14.01\text{g N}}{1\text{mol N}}\right) = 28.02\text{g N};
\text{mass}\% \text{ N} &= \left(\frac{28.02}{96.09}\right) \times 100\% = 29.16\%.
\end{align*}
\]

6. d \(\{\text{since there is “more” O in CO}_3^{-2} \text{ than in CO or CO}_2, \text{ CO and CO}_2 \text{ can be eliminated; now determine the %O in H}_2\text{O and CO}_3^{-2} \text{ with two calculations; for H}_2\text{O: mass}\% \text{ O} = \text{mass O/total mass x 100%} = 16/18 \times 100\% = 89\%; \text{ for CO}_3^{-2}: \text{mass}\% \text{ O} = \text{mass O/total mass x 100%} = 48/60 \times 100\% = 80\%}\}

7. \(\text{C}_3\text{H}_5\text{O}_2 \text{ is the empirical formula; C}_6\text{H}_{10}\text{O}_4 \text{ is the molecular formula} \quad \{\text{assume 100g;}\}

\[
\begin{align*}
49.3 \text{g C} \left(\frac{1\text{mol C}}{12.01\text{g C}}\right) &= 4.105\text{mol C}; \quad 6.90\text{g H} \left(\frac{1\text{mol H}}{1.008\text{g H}}\right) = 6.845\text{mol H}; \quad 43.8\text{g O} \left(\frac{1\text{mol O}}{16.00\text{g O}}\right) = 2.738\text{mol O};
\text{C}_4\text{H}_{10}\text{O}_5\text{O}_2 \text{; divide by 2.738:} \quad \text{C}_4\text{H}_{10}\text{O}_5\text{O}_2 \left(\frac{2.738}{2.738}\right) = \text{C}_4\text{H}_{10}\text{O}_5\text{O}_2 \text{; fraction of 0.5} \rightarrow 1/2 \rightarrow x2:\n\text{C}_4\text{H}_{10}\text{O}_5\text{O}_2 \left(\frac{1.499 \times 2}{2.738}\right) \left(\frac{2.500 \times 2}{2.738}\right) \left(\frac{1 \times 2}{2.738}\right) = \text{C}_3\text{H}_5\text{O}_2 = \text{EF; molar mass EF} = 73.07\text{g/mol}; \quad \text{ratio} = \frac{\text{molar mass molecular formula}}{\text{molar mass empirical formula}} = \frac{146}{73.07} = 1.998 = 2; \quad \text{multiply EF by 2; C}_3\text{H}_5\text{O}_2 \left(\frac{3 \times 2}{2.738}\right) \left(\frac{5 \times 2}{2.738}\right) \left(\frac{2 \times 2}{2.738}\right) = \text{C}_6\text{H}_{10}\text{O}_4 = \text{MF}\}\]

8. Ca \(\{\text{Need to determine average atomic mass (AAM) of M; to do this you need g M and mol M;}\)

\[
\begin{align*}
g_{\text{total}} &= \text{g M} + \text{g Cl} \rightarrow \text{g M} = 15.00 - 9.583 = 5.417\text{g M};
\text{mol M:} \quad 9.583\text{g Cl} \left(\frac{1\text{mol Cl}_2}{70.90\text{g Cl}_2}\right) \left(\frac{2\text{mol Cl}}{1\text{mol Cl}_2}\right) \left(\frac{1\text{mol M}}{2\text{mol Cl}}\right) = 0.1352\text{mol M};
\text{molar mass M} = \frac{\text{grams M}}{\text{mol M}} = \frac{5.417\text{g M}}{0.1352\text{mol M}} = 40.07\text{g/mol} \rightarrow \text{Ca from the periodic table}\}\]

9. \(107.9\text{g/mol} \quad \{\text{Need to determine molar mass of M; to do this you need g M and mol M; assume 100g} \rightarrow 93.10\text{gM; if 93.10\% is M, then 100.00\%=93.10\%} = 6.90\% \rightarrow \text{6.90g O; }\)

\[
\text{6.90g O} \left(\frac{1\text{mol O}}{16.00\text{g O}}\right) \left(\frac{2\text{mol M}}{1\text{mol O}}\right) = 0.8625\text{mol M}; \quad \text{molar mass M} = \frac{\text{grams M}}{\text{mol M}} = \frac{93.10\text{g M}}{0.8625\text{mol M}} = 107.94\text{g/mol}\}\]