WKS 3-11 – Unit 3 Review – Chem Honors
Moles & Stoichiometry

NAME Answer Key
Period ______ **Date** ________________

Material covered on test:

- **Labs:** ~~Percent Composition/Empirical Formula Lab~~, $Cu_3(PO_4)_2$ Lab, Micro-Mole Rockets
- **Coverage in Book:** Empirical and Molecular Formulas: p 229-233;
  Mole Conversions: p. 78-85, p221-233
  Balancing Equations, p.241-253; Stoichiometry p.275-294

TEST Format: 70-80 pts total: multiple choice, short explanations, calculations, lab type questions (including sample data)

Topics covered on test:

- History of moles: Law of conservation of mass, Law of Definite proportions and Law of multiple proportions; Theories of Dalton, Gay-Lussac, and Avogadro
- Avogadro's Hypothesis: How it's used to balance equations (& determine formulas) and relative masses of gases
- Mole conversions (Mole Road) will not be tested explicitly but may be part of a larger problem.
- Percent composition, empirical formulas and molecular formulas
- Balancing equations
- Stoichiometry problems: Conversions between amounts of one substance to amounts of another (in a reaction)
- Limiting reactants, theoretical yield and percent yield
- Lab type questions-- general lab separation techniques, procedures, concepts and calculations. You may be given sample data to analyze

I. Practice Questions

1) Compare and contrast the Law of Definite Proportions to the Law of Multiple Proportions.
   The Law of Definite Proportions describes the constant composition of any two samples of the same compound (e.g. H_2O is always 2 parts H and 1 part O, or by mass 1 g H to 8 g O). The Law of Multiple Proportions also describes the compositions of compounds, but compares the different ratios of the same elements in two different compounds (e.g. H_2O vs. H_2O_2 or CO_2 vs. CO).

2) What did Gay-Lussac contribute to our understanding of reactions of gases?
   He showed that gases react in simple whole number ratios of volumes. (He did lots of reactions with gases and measured the volumes involved. He collected the data, but he did not explain why these results are obtained.)

3) 1 mole of O_2 gas has the same volume of 1 mole of CH_4 gas (at same T + P). However, 1 mole of liquid O_2 does not have the same volume as 1 mole of liquid CH_4. Explain why. (*Hint: think about the sizes of the individual molecules.)*
   In gases, the particles are very far apart, and the volume occupied by the sample is large compared to the size of each particle, so their size is inconsequential. In a liquid, however, the particles are in contact with each other, so the volume occupied by the sample depends on the size of the particles.

4) What is the difference between an empirical formula and a molecular formula?
   An empirical formula is the simplest, most reduced whole-number ratio of the elements in a compound; a molecular formula is the exact, unreduced formula showing the actual numbers of atoms of each element in the compound.

5) What is the empirical formula for the substance with the molecular formula of $C_2H_4Br_2$? CH_2Br

6) As shown below, one liter of phosphorous vapor will exactly react with 6 liters of chlorine vapor to produce 4 liters of phosphorous trichloride vapor.

Volume ratio: 1 L of phosphorous vapor + 6 L chlorine vapor → 4 L phosphorous trichloride vapor

$\underline{\ 1\ }\ P_{\boxed{4}}\ (g) + \underline{\ 6\ }\ Cl_{\boxed{2}}\ (g) \rightarrow \underline{\ 4\ }\ PCl_3\ (g)$

Particle Ratio: 1 particle 6 particles 4 particles

a) Determine the coefficients in the chemical equation by filling in the "blanks" above. Explain your reasoning by discussing how you used a particular scientific concept to determine these coefficients.

- Since the chlorine vapor has a volume 6 times that of the phosphorous vapor, the chlorine vapor has 6 times the number of particles. This logic uses Avogadro's hypothesis, because the number of particles is proportional to the volumes (at same T & P), so 6 times the volume has 6 times the particles.
- Likewise, the PCl_3 has 4 times the volume of the phosphorus vapor, so 1 particle of phosphorous produces 4 particles of PCl_3. **The coefficients represent the particle ratios.**

b) Determine the subscripts needed in the phosphorous formula and the chlorine formula by filling the "boxes." Explain your reasoning by *discussing* the scientific concept used to determine these subscripts, then explain how you derived the subscripts in this specific example.

- *The subscripts are determined by equalizing the number of each type of atom on both sides of the equation. This holds true because of the Law of Conservation of Mass. The mass stays constant throughout a chemical reaction because no atoms are created or destroyed—just rearranged.*
- *In this equation, there are 4 atoms of P on the right. Thus, there must be 4 atoms of P on the left, but all 4 atoms must be in one particle. Thus, there is one particle of P_4. Likewise, there are 12 atoms of Cl on the right. Thus, there must be 12 atoms of Cl on the left, but in 6 particles. Thus, there are 6 particles of Cl_2.*

7) Balance the following equations using the *lowest whole number* coefficients:

a) ____ $Fe_2(SO_4)_3$ + __3__ $Ba(NO_3)_2$ → __2__ $Fe(NO_3)_3$ + __3__ $BaSO_4$

b) __2__ C_3H_8O + __9__ O_2 → __8__ H_2O + __6__ CO_2

c) __3__ $Ca(OH)_2$ + __2__ H_3PO_4 → ____ $Ca_3(PO_4)_2$ + __6__ H_2O

II. Moles Calculations

8) Aluminum has a density of 2.71 g/cm³. What is the volume of one Al atom? *(What is the mass of 1 atom?)*

$$1\text{ atom Al}\times\frac{1\text{ mol Al}}{6.022\times10^{23}\text{ atoms Al}}\times\frac{26.98\text{ g Al}}{1\text{ mol Al}}\times\frac{1\text{ cm}^3}{2.71\text{ g}}=\boxed{1.65\times10^{-23}\text{ cm}^3}$$

9) An unknown sample contains only C & H. If 26.8 g of the sample contains 4.90 g H, what is its % composition?

$$\text{Mass}_C = 26.8\text{ g} - 4.90\text{ g} = 21.9\text{ g C}$$

$$\%C=\frac{21.9\text{ g}}{26.8\text{ g}}\times100\%=\boxed{81.7\%\text{ C}};\%H=\frac{4.90\text{ g}}{26.8\text{ g}}\times100\%=\boxed{18.3\%\text{ H}}$$

10) Octane is a hydrocarbon that is an important component in gasoline.

a) Octane contains 84.12% carbon and 15.88% hydrogen (all by mass). What is its empirical formula?

$$\left.\begin{array}{l}84.12\text{ g C}\times\dfrac{1\text{ mol C}}{12.01\text{ mol C}}=7.004\text{ mol C}\div7.004=1\times4=4\\[2ex]15.88\text{ g H}\times\dfrac{1\text{ mol H}}{1.008\text{ g H}}=15.75\text{ mol H}\div7.004=2.249\times4=9\end{array}\right\}\text{EF}=C_4H_9$$

b) The molar mass of octane is 114 g/mol. What is its molecular formula?

$$\text{EF Mass}=4(12.01)+9(1.008)=57.11\text{ g/mol}$$

$$\text{Multiplier n}=\frac{114\text{ g/mol}}{57.11\text{ g/mol}}=2;\ \text{Molecular Formula}=(C_4H_9)\times2=\boxed{C_8H_{18}}$$

11) What is the percent composition (by mass) of $Sr_3(PO_4)_2$? (all elements!)

$MM = 3(87.62) + 2(30.97) + 8(16.00) = 452.8 \text{ g/mol}$

$\%Sr = \frac{262.86 \text{ g}}{452.8 \text{ g}} \times 100\% = \boxed{58.05\% \text{ Sr}}$; $\%P = \frac{61.94 \text{ g}}{452.8 \text{ g}} \times 100\% = \boxed{13.68\% \text{ P}}$; $\%O = \frac{128.0 \text{ g}}{452.8 \text{ g}} \times 100\% = \boxed{28.27\% \text{ O}}$

12) A student performed an experiment to determine the mass % of sulfur in Na_2SO_3. She combined 18.36 g Na_2SO_3 with 10.24 g of HCl solution and observed the following reaction:

$$Na_2SO_3 \text{ (s)} + 2 \text{ HCl (aq)} \rightarrow 2 \text{ NaCl (aq)} + H_2O \text{ }(l) + SO_2 \text{ (g)}$$

Bubbles were observed, indicating formation of the SO_2. When the reaction was complete, the total mass of the mixture was 19.27 g and all of the mass lost was due to loss of the SO_2 gas, which contained all of the S. Find the mass of S in the SO_2 and then use that mass to determine the mass % of S in Na_2SO_3.

$\text{mass } SO_2 = 18.36 \text{ g} + 10.24 \text{ g} - 19.27 \text{ g} = 9.33 \text{ g } SO_2$; $MM_{SO_2} = 32.066 \text{ g} + 2(16.000 \text{ g}) = 64.066 \text{ g/mol}$

$$\text{mass S} = 9.33 \text{ g } SO_2 \times \frac{1 \text{ mol } SO_2}{64.066 \text{ g } SO_2} \times \frac{1 \text{ mol S}}{1 \text{ mol } SO_2} \times \frac{32.066 \text{ g S}}{1 \text{ mol S}} = 4.67 \text{ g S}$$

$$\%S = \frac{4.67 \text{ g S}}{18.36 \text{ g } Na_2SO_3} \times 100\% = \boxed{25.4\% \text{ S}}$$

III. Stoichiometry Calculations

13) Liquid octane (C_8H_{18}) burns in oxygen according to this unbalanced equation (balance it):

$$\underline{\ 2\ } \ C_8H_{18} \text{ (l)} + \underline{\ 25\ } \ O_2 \text{ (g)} \rightarrow \underline{\ 16\ } \ CO_2 \text{ (g)} + \underline{\ 18\ } \ H_2O \text{ (g)}$$

How many liters of octane can be burned with 68,500 L oxygen at STP? *Density of octane = 0.702 g/mL*

$MM_{C_8H_{18}} = 8(12.01) + 18(1.008) = 114.2 \text{ g/mol}$

$$68{,}500 \text{ L } O_2 \times \frac{1 \text{ mol } O_2}{22.4 \text{ L } O_2} \times \frac{2 \text{ mol } C_8H_{18}}{25 \text{ mol } O_2} \times \frac{114.2 \text{ g } C_8H_{18}}{1 \text{ mol } C_8H_{18}} \times \frac{1 \text{ mL } C_8H_{18}}{0.702 \text{ g } C_8H_{18}} \times \frac{1 \times 10^{-3} \text{ L}}{1 \text{ mL}} = \boxed{39.8 \text{ L } C_8H_{18}}$$

14) This reaction is carried out: ____ P_4 (s) + $\underline{\ 6\ }$ $Cl_2 \rightarrow$ $\underline{\ 4\ }$ PCl_3 (balance!)

0.130 moles of P_4 is reacted with 1.392 moles of Cl_2. What is the limiting reactant? Show a calc and a sentence to explain.

$$\frac{0.130 \text{ mol } P_4}{1 \text{ mol } P_4} = 0.130 \text{ reactions}; \quad \frac{1.392 \text{ mol } Cl_2}{6 \text{ mol } Cl_2} = 0.232 \text{ reactions}$$

P_4 is limiting since it makes fewer reactions than Cl_2.

15) Hydrofluoric acid solutions cannot be stored in glass containers because HF reacts readily with silica (SiO_2) in glass to produce hexafluorosilicic acid (H_2SiF_6): $SiO_2\text{(s)} + 6 \text{ HF(aq)} \rightarrow H_2SiF_6\text{(aq)} + 2 \text{ } H_2O(l)$

Suppose 45.2 g SiO_2 and 88.2 g HF are mixed together and 91.5 g H_2SiF_6 are produced.

a) What is the theoretical yield of H_2SiF_6 (in grams)?

$$45.2 \text{ g } SiO_2 \times \frac{1 \text{ mol } SiO_2}{60.08 \text{ g } SiO_2} = 0.752 \text{ mol } SiO_2; \quad \frac{0.752 \text{ mol } SiO_2}{1 \text{ mol } SiO_2} = 0.752 \text{ reactions}$$

$$88.2 \text{ g HF} \times \frac{1 \text{ mol HF}}{20.01 \text{ g HF}} = 4.41 \text{ mol HF}; \quad \frac{4.41 \text{ mol HF}}{6 \text{ mol HF}} = 0.735 \text{ reactions}$$ *HF is Limiting Reactant*

$$4.41 \text{ mol HF} \times \frac{1 \text{ mol } H_2SiF_6}{6 \text{ mol HF}} \times \frac{144.1 \text{ g } H_2SiF_6}{1 \text{ mol } H_2SiF_6} = 105.9 \text{ g } H_2SiF_6 = \boxed{106 \text{ g } H_2SiF_6}$$

b) What is the limiting reactant? Explain how you know with one sentence.
HF is the limiting react. According to the calculations, only 0.735 reactions can be obtained with HF compared to 0.752 reactions with SiO_2.

c) How many grams of the excess reactant are left unreacted?

$$\text{g SiO}_2\text{ used} = 4.41\text{ mol HF} \times \frac{1\text{ mol SiO}_2}{6\text{ mol HF}} \times \frac{60.08\text{ g SiO}_2}{1\text{ mol SiO}_2} = 44.14\text{ g SiO}_2 = \boxed{44.1\text{ g SiO}_2\text{ used}}$$

$$\text{g SiO}_2\text{ remaining} = 45.2\text{ g SiO}_2 - 44.1\text{ g SiO}_2 = \boxed{1.1\text{ g SiO}_2\text{ remaining}}$$

d) What is the percent yield of H_2SiF_6?

$$\%\text{ Yield} = \frac{91.5\text{ g H}_2\text{SiF}_6}{106\text{ g H}_2\text{SiF}_6} \times 100\% = \boxed{86.3\ \%}$$

IV. Extra Review—More practice of the more difficult types of calculations

16) Suppose 6.00 L of hydrogen gas is reacted with excess nitrogen gas at standard temperature and pressure to form ammonia by the following balanced chemical equation: N_2 (g) + 3 H_2 (g) → 2 NH_3 (g)
If the ammonia gas produced is cooled until it liquefies, what volume of liquid ammonia should be collected? *Density of liquid ammonia = 0.674 g/mL*

$$6.00\text{ L H}_2 \times \frac{1\text{ mol H}_2}{22.4\text{ L H}_2} \times \frac{2\text{ mol NH}_3}{3\text{ mol H}_2} \times \frac{17.03\text{ g NH}_3}{1\text{ mol NH}_3} \times \frac{1\text{ mL NH}_3}{0.674\text{ g NH}_3} = 4.512\text{ mL NH}_3 = \boxed{4.51\text{ mL NH}_3}$$

17) Hydrogen is generated by passing hot steam of iron, which oxidizes to form Fe_3O_4, in the following *unbalanced* equation (balance it first!):

_3_ Fe (s) + _4_ H_2O (g) → _4_ H_2 (g) + ___ Fe_3O_4 (s)

Suppose 9.78 L of hydrogen gas is produced at STP when 21.5 g Fe reacts with 15.6 g H_2O.

a) What is the theoretical yield of hydrogen gas (in L, at STP)?

$$21.5\text{ g Fe} \times \frac{1\text{ mol Fe}}{55.85\text{ g Fe}} = 0.385\text{ mol Fe};\ \frac{0.385\text{ mol Fe}}{3\text{ mol Fe}} = 0.128\text{ reactions *Fe Limiting Reactant*}$$

$$15.6\text{ g H}_2\text{O} \times \frac{1\text{ mol H}_2\text{O}}{18.02\text{ g H}_2\text{O}} = 0.866\text{ mol H}_2\text{O};\ \frac{0.866\text{ mol H}_2\text{O}}{4\text{ mol H}_2\text{O}} = 0.216\text{ reactions}$$

$$0.385\text{ mol Fe} \times \frac{4\text{ mol H}_2}{3\text{ mol Fe}} \times \frac{22.4\text{ L H}_2}{1\text{ mol H}_2} = 11.497\text{ L H}_2 = \boxed{11.5\text{ L H}_2}$$

b) What is the percent yield of hydrogen gas?

$$\%\text{ Yield} = \frac{9.78\text{ L H}_2}{11.5\text{ L H}_2} \times 100\% = \boxed{85.0\ \%\text{ H}_2}$$

c) How many grams of the excess reactant are left over?

$$\text{g H}_2\text{O used} = 0.385\text{ mol Fe} \times \frac{4\text{ mol H}_2\text{O}}{3\text{ mol Fe}} \times \frac{18.02\text{ g H}_2\text{O}}{1\text{ mol H}_2\text{O}} = 9.249\text{ g H}_2\text{O} = 9.25\text{ g H}_2\text{O used}$$

$$\text{g H}_2\text{O remaining} = 15.6\text{ g H}_2\text{O} - 9.25\text{ g H}_2\text{O} = 6.35\text{ g H}_2\text{O} = \boxed{6.4\text{ g H}_2\text{O remaining}}$$