PRACTICE EXAM #1

Chem1A, General Chemistry I

1.) Isopropanol, commonly known as rubbing alcohol, has a chemical formula of C_3H_7OH.

a.) Write the **balanced** chemical equation for the combustion of isopropanol.

$C_3H_7OH(l) + 9/2\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$

b.) Calculate the grams of carbon dioxide produced when 345 mL of C_3H_7OH combusts, provided that the density of C_3H_7OH is 0.786 g/cm^3.

Use dimensional analysis and stoichiometry to convert 345 mL C_3H_7OH to g CO_2.

$$345\text{ mL } C_3H_7OH \times \frac{1\text{ cm}^3}{1\text{ mL}} \times \frac{0.786\text{ g } C_3H_7OH}{1\text{ cm}^3} \times \frac{1\text{ mol } C_3H_7OH}{60.094\text{ g } C_3H_7OH} \times \frac{3\text{ mols } CO_2}{1\text{ mol } C_3H_7OH} \times \frac{44.01\text{ g } CO_2}{1\text{ mol } CO_2}$$

$$= 59\underline{5}.7762023 \rightarrow \boxed{\mathbf{596\text{ g } CO_2}}$$

2.) An unknown element is a reddish liquid at room temperature and has a mass number of 80. If the element carries an anionic charge of 1, how many protons, neutrons, and electrons does one atom of this element contain?

Reddish liquid – bromine (Z = 35, so $p^+ = 35$)

$p^+ =$ **35** ; $e^- = (p^+ + 1 = 35 + 1) =$ **36**, $n = (A - p^+ = 80 - 35) =$ **45**.

3.) Name the following compounds appropriately.

a.) Cs_2SO_4 cesium sulfate

b.) I_2Cl_6 diiodine hexachloride

c.) $HNO_2(aq)$ nitrous acid

d.) $Cu(OH)_2$ copper(II) hydroxide

4.) Give the balanced ionic or molecular formulas for the following compounds.

a.) perchloric acid $HClO_4$

b.) tetrasulfur dioxide S_4O_2

c.) zinc phosphate $Zn_3(PO_4)_2$

d.) manganese(III) carbonate $Mn_2(CO_3)_3$

5.) A compound contains 10.7% C, 46.4% Cr, and the rest as O by mass. It has a molecular weight of 112.0 g/mol. Determine its molecular formula and name it appropriately. (25 pts)

Calculate the % O by subtracting % C and % Cr from 100%.

100% - 10.7% - 46.4% = 42.$\underline{9}$% O

Assume 100 g of total sample to convert mass percent to grams. Convert grams of element to mols and divide by the smallest.

$10.7\text{ g C} \times \frac{1\text{ mol C}}{12.01\text{ g C}} = 0.89\underline{0}9242298\text{ mols C}/0.891\text{ mols} \rightarrow 1\text{ mol C}$

$46.4\text{ g Cr} \times \frac{1\text{ mol Cr}}{52.00\text{ g Cr}} = 0.89\underline{2}3076923\text{ mols Cr}/0.891\text{ mols} \rightarrow 1.00\text{ mols Cr}$

$42.9\text{ g O} \times \frac{1\text{ mol C}}{16.00\text{ g C}} = 2.6\underline{8}125\text{ mols O}/\ 0.891\text{ mols} \rightarrow 3.00\text{ mols O}$

Empirical formula: $CrCO_3$

Empirical weight = (1)(52.00 g/mol) + (1)(12.01 g/mol) + (3)(16.00 g/mol) = 112.0$\underline{1}$ g/mol

(112.0 g/mol)/(112.0$\underline{1}$ g/mol) = 1 so $CrCO_3$ is the molecular formula

$\boxed{\mathbf{CrCO_3 \rightarrow \text{chromium(II) carbonate}}}$

6.) Boron has two main isotopes: ^{10}B and ^{11}B. The ^{10}B isotope has a mass of 10.012 amu while the ^{11}B isotope is 11.009 amu, and an atomic mass of 10.811 amu. Determine the natural abundance of each isotope.

% ^{10}B + % ^{11}B = 100% since only two isotopes. Make % ^{10}B = x. So then, % ^{11}B = 1-x.

10.811 amu = (x)(10.012 amu) + (1 – x)(11.009 amu)

10.811 amu = (10.012 amu)(x) + 11.009 amu – (11.009 amu)(x)

10.811 amu – 11.009 amu = (10.012 amu)(x) – (11.009 amu)(x)

-0.19$\underline{8}$ amu =(-0.99$\underline{7}$ amu)(x)

x = 0.19$\underline{8}$5957874 x 100% so then

$\boxed{\mathbf{\%\ ^{10}B = 19.9\% \text{ and } \%\ ^{11}B = 80.1\%}}$

7.) Chlorine gas will combine with solid phosphorous (P_4) to synthesize solid phosphorous trichloride.

a.) Write the **balanced equation** for this reaction.

$6\ Cl_2(g) + P_4(s) \rightarrow 4\ PCl_3(s)$

b.) Determine the limiting reactant and theoretical yield, in g, when 15.86 g of phosphorous react with 23.59 g chlorine.

Convert both reactants to maximum amount of product. Smallest produced is limiting reactant/theoretical yield.

$$15.86\text{ g } P_4 \times \frac{1\text{ mol } P_4}{123.88\text{ g } P_4} \times \frac{4\text{ mols } PCl_3}{1\text{ mol } P_4} \times \frac{137.32\text{ g } PCl_3}{1\text{ mol } PCl_3} = 70.3\underline{2}273813\text{ g } PCl_3$$

$$23.59\text{ g } Cl_2 \times \frac{1\text{ mol } Cl_2}{70.90\text{ g } Cl_2} \times \frac{4\text{ mols } PCl_3}{6\text{ mols } Cl_2} \times \frac{137.32\text{ g } PCl_3}{1\text{ mol } PCl_3} = 30.4\underline{5}96032\text{ g } PCl_3$$

Theoretical Yield: 30.46 g PCl_3, limiting reactant is Cl_2

c.) Calculate the amount of reactant in excess left over when the reaction is complete, in g.

Convert the theoretical yield of product back to reactant in excess to calculate how much is used. Take the difference between total amount – used amount.

$$30.46\text{ g } PCl_3 \times \frac{1\text{ mol } PCl_3}{137.32\text{ g } PCl_3} \times \frac{1\text{ mol } P_4}{4\text{ mols } PCl_3} \times \frac{123.88\text{ g } P_4}{1\text{ mol } P_4} = 6.86\underline{9}692689\text{ g } P_4\text{ used}$$

$15.86\text{ g } P_4$ total – $6.87\underline{0}\text{ g } P_4$ used $= 8.9\underline{9}0307311\text{ g} \rightarrow$ **8.99 g P_4 in excess**

d.) If the percent yield for this reaction was 85.67%, calculate the **actual yield** of product, in g.

Use the equation for percent yield and solve for actual yield (x).

$$85.67\ \% = \frac{x}{30.46\text{ g } PCl_3} \times 100\%$$

$x = (0.8567)(30.46\text{ g } PCl_3) = 26.0\underline{9}5082 \rightarrow$ **26.10 g PCl_3**