Name _____ **AP Chemistry** ___/___/___

Chapter 3 Practice Problems Homework

1. There are two naturally occurring isotopes of Gallium. One isotope, gallium-69, has a mass of 68.926 amu and an abundance of 60.108%. Calculate the **mass and name** of the other stable isotope of gallium.

$69.72 = (0.60108)(68.926) + (0.39892)(x)$

$6972 = 41.430 + 0.39892x$

$28.29 = 0.39892x$

$x = 70.92$ amu

Gallium-71

2. A sample of xenon fluoride contains molecules with the general formula XeF_n, where n is some whole number. If 9.03×10^{20} molecules of XeF_n weigh 0.368 g, determine the formula of the sample.

$9.03 \times 10^{20} \div 6.022 \times 10^{23} = 0.00150$ mole

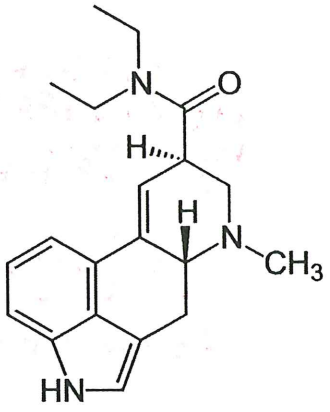
$0.368 \div 0.00150 = 245.4$

$245.4 - 131.29 = 114$ g F_n

$114 \div 19 = 6$

XeF_6

3. D-lysergic acid diethylamide, (LSD) has a chemical formula of: $C_{20}H_{25}N_3O$.



a. Calculate the molar mass of LSD. Use the full atomic weight as it is listed on your periodic table.

C: $20 \times 12.01 = 240.20$

H: $25 \times 1.008 = 25.20$

N: $3 \times 14.01 = 42.03$

O: $1 \times 16.00 = 16.00$

323.43 amu or g/mol

b. Calculate the percent composition of the elements in LSD.

C: $240.20 \div 323.43 \times 100 = 74.226\%$

H: $25.20 \div 323.43 \times 100 = 7.791\%$

N: $42.03 \div 323.43 \times 100 = 13.00\%$

O: $16.00 \div 323.43 \times 100 = 4.97\%$

c. Determine the mass of each element in a 297 milligram sample of LSD.

This amount killed an elephant.

C: $0.74226 \times 297 = 221$ mg

H: $0.07791 \times 297 = 23.1$ mg

N: $0.1300 \times 297 = 38.6$ mg

O: $0.0497 \times 297 = 14.7$ mg

d. In the 1960's a standard dose of LSD contained 270 micrograms of LSD. How many molecules of LSD are the in this size dose?

$270 \times 10^{-6} \div 323.43 \times 6.022 \times 10^{23} = 5.0 \times 10^{17}$

e. What is the mass of 1 molecule of LSD?

$1 \div 6.022 \times 10^{23} \times 323.43 = 5.37 \times 10^{-22}$ g

4. A compound containing the elements C, H, N and O is analyzed. When a 2.1106 g sample is burned in excess oxygen, 3.2017 g of $CO_2(g)$ is formed. The combustion analysis also showed that the sample contained 0.1710 g of hydrogen.

(i) Determine the mass, in grams, of C in the 2.1106 g sample of the compound.

$3.2017 \times \frac{12.01}{44.01} = 0.8737\text{g}$

(ii) When the compound is analyzed for N content only, the mass percent of N is found to be 32.16%. Determine the mass, in grams of N in the original 2.1106 g sample of the compound.

$0.3216 \times 2.1106 = 0.6788$

(iii) Determine the mass, in grams, of oxygen in the original 2.1106 g sample of the compound.

C N H

$2.1106 - 0.8737 - 0.6788 - 0.1710 = 0.3871$

(iv) Determine the empirical formula of the compound.

C: $0.8737 \div 12.01 = 0.07275$
N: $0.6788 \div 14.01 = 0.04845$
H: $0.1710 \div 1.008 = 0.1696$
O: $0.3871 \div 16.00 = 0.02419$

$\div 0.02419 = 3, 2, 7, 1$ $C_3N_2H_7O$

(v) The molecular mass of the compound is 174.2 g/mol. Determine the molecular formula of the compound.

$C_3N_2H_7O$

$36 + 28 + 7 + 16 = 87$

$\frac{174.2}{87} = 2$ $C_6N_4H_{14}O_2$

5. Nitrogen gas reacts with hydrogen gas to form gaseous ammonia.

a. Write a balanced chemical equation for this reaction.

$N_2 + 3H_2 \rightarrow 2NH_3$

b. If 65.0 grams of nitrogen react with 10.0 grams of hydrogen gas, what is your limiting reactant. Justify your answer by showing your work or reasoning.

$65.0\text{g } N_2 \div 28.02 = 2.32 \text{ mol } N_2$
$10.0\text{g } H_2 \div 2.02 = 4.95 \text{ mol } H_2$

H_2 is the Limiting Reactant b/c you don't have 3 times as much.

c. How many grams of the excess reagent will be left over once the reaction in part B has been carried out?

$10.0\text{g } H_2 \left| \frac{1\text{mol } H_2}{2.016\text{g } H_2} \right| \frac{1\text{mol } N_2}{3\text{mol } H_2} \left| \frac{28.02\text{g } N_2}{1\text{mol } N_2} \right| = 46.3\text{g } N_2 \text{ needed}$

$65.0 - 46.3\text{g} = 18.7\text{g excess } N_2$

d. If in using the above amounts from part B, what is your theoretical yield?

$10.0\text{g } H_2 \left| \frac{1\text{mol } H_2}{2.016\text{g } H_2} \right| \frac{2\text{mol } NH_3}{3\text{mol } H_2} \left| \frac{17.03\text{g } NH_3}{1\text{mol } NH_3} \right| = 56.3\text{g } NH_3 \text{ produced}$

e. If in using the above amounts from part B you produce 32.0 grams of ammonia, what is your percent yield?

$\frac{32.0}{56.3} \times 100 = 56.8\%$