Chemistry 130
Stoichiometry Practice 2

1) What is the weight percent of nitrogen in N_2O_5?

2) What mass of N_2O_5 contains 25.0 g of N? Do this in 2 ways, one using weight percent and one using stoichiometry (moles).

3) What is the weight percent of ammonium in $(NH_4)_2SO_4$?

4) When 1.000 g of copper was burned in an excess of O_2 gas, a compound with mass 1.126 g was formed. What is the empirical formula of the compound?

5) An oxide of nitrogen is 34.44 % N by mass. What is the compound's empirical formula?

6) A compound is 55.80 % C, 7.03 % H, and 37.17 % O by mass. What is the compound's empirical formula?

7) For a hypothetical compound between elements A and B, a student obtained the crude formula $A_{1.60}B$. What is the correct empirical formula for the compound?

8) A compound with empirical formula $C_4H_5N_2O$ was found to have a molecular weight of around 190 g/mol. What is the compound's molecular formula?

9) a) A lab student decomposed 1.241 g of $AlCl_3$ and was left with 0.262 g of Al. From the experimental data, what mass of Cl was in the original sample?
   b) According to the student's data, what is the experimental weight percent of Cl in $AlCl_3$?
   c) According to the chemical formula for $AlCl_3$, what is the actual (true) weight percent of Cl in $AlCl_3$?
   d) To the correct significant figures, what is the percent error of the student's weight percent of Cl?

10) Balance the equation for the combustion of C_3H_7OH

The next 5 questions refer to the following balanced equation:

$$Mg_3N_2(s) + 6\ H_2O(l) \rightarrow 3\ Mg(OH)_2(s) + 2\ NH_3(aq)$$

11) What mass of ammonia can be formed by reacting 10.0 g of Mg_3N_2?

12) What mass of water is necessary to produce 4.5×10^{25} magnesium hydroxide formula units?

13) What mass of magnesium hydroxide can be formed by reacting 20.0 g of H_2O with 10.0 g of Mg_3N_2?

14) What is the mass of excess reagent (i.e. reactant) leftover in Question 13? To do this, first use the moles of limiting reactant to figure out how many moles of excess reactant are used; from there you can figure out how many moles (and thus what mass) of excess reactant are left.

15) What is the percent yield of the reaction if only 10.0 g of magnesium hydroxide were collected from the reaction in Question 13?

16) The next 2 questions refer to the balanced reaction **$5\ O_2(g) + 2\ NiS_2(s) \rightarrow 2\ NiO(s) + 4\ SO_2(g)$**

   a) Calculate the theoretical yield of SO_2 if 9.86 g of O_2 are reacted with 11.20 g of NiS_2.

   b) If the actual yield of SO_2 is 10.00 g, what is percent yield of the reaction?

Solutions

1) What is the weight percent of nitrogen in N_2O_5?

The molar mass of N_2O_5 is 2(14.01) + 5(16.00) = 108.02 g/mol

So mass percent N = $\frac{2(14.01)}{108.02} \times 100 = 25.94\ \%$

2) What mass of N_2O_5 contains 25.0 g of N? Do this in 2 ways, one using weight percent and one using stoichiometry (moles).

Weight percent: $\frac{P}{W} \times 100 = \%$ → $\frac{25.0\ g}{W} \times 100 = 25.94$ Solve to get W = 96.4 g

Stoichiometry:

$$\frac{25.0\ g\ N}{14.01\ g/mol} = 1.78444\ mol\ N \times \frac{1\ mol N_2O_5}{2\ mol\ N} \times \frac{108.02\ g}{mol\ N_2O_5} = 96.4\ g\ N_2O_5$$

3) What is the weight percent of ammonium in $(NH_4)_2SO_4$?

The molar mass of the NH_4 unit is 14.01 + 4(1.01) = 18.05 g/mol; the M.M. of $(NH_4)_2SO_4$ is 132.16 g/mol.

Therefore we have $\frac{2(18.05)}{132.16} \times 100 = 27.32\ \%$ ammonium by mass

4) When 1.000 g of copper was burned in an excess of O_2 gas, a compound with mass 1.126 g was formed. What is the empirical formula of the compound?

Mass of Cu in compound = 1.000 g; mass of O in compound = 1.126 g – 1.000 g = 0.126 g

Therefore $\frac{1.000\ g}{63.54\ g/mol} = 0.01574\ mol\ Cu$ and $\frac{0.126\ g}{16.00\ g/mol} = 0.00788\ mol\ O$

So the compound is $Cu_{0.01574}O_{0.00788}$

Divide both subscripts by the smallest number (0.00788) → Cu_2O is the empirical formula

5) An oxide of nitrogen is 34.44 % N by mass. What is the compound's empirical formula?

Assume 100.00 g of compound

$\frac{34.44\ g\ N}{14.01\ g/mol} = 2.458\ mol\ N$ and $\frac{65.56\ g\ O}{16.00\ g/mol} = 4.098\ mol\ O$

So compound is $N_{2.458}O_{4.098}$

Divide both subscripts by smallest number (2.458) → $N\ O_{1.667}$

Multiply both coefficients by successive integers until both coefficients are integers

→ $N_{1x3}O_{1.667x3} = \boxed{N_3O_5}$

6) A compound is 55.80 % C, 7.03 % H, and 37.17 % O by mass. What is the compound's empirical formula?

Assume 100.00 g of compound

$\frac{55.80\ g\ C}{12.01\ g/mol} = 4.646\ mol\ C$ $\frac{7.03\ g\ H}{1.01\ g/mol} = 6.96\ mol\ H$ $\frac{37.17\ g\ O}{16.00\ g/mol} = 2.323\ mol\ O$

So the compound is $C_{4.646}H_{6.96}O_{2.323}$

Divide all subscripts by the smallest mole number (2.323) → C_2H_3O

7) For a hypothetical compound between elements A and B, a student obtained the crude formula $A_{1.60}B$. What is the correct empirical formula for the compound?

Start multiplying 1.60 by integers until the result is within 0.1 of a whole number….

1.60 x 2 = 3.20 (no)
1.60 x 3 = 4.80 (no)
1.60 x 4 = 6.40 (no)
1.60 x 5 = 8.00 (yes!) → So we have $A_{(1.60x5)}B_{(1x5)}$ and the correct empiricial formula is $\boxed{A_8B_5}$

8) A compound with empirical formula $C_4H_5N_2O$ was found to have a molecular weight of around 190 g/mol. What is the compound's molecular formula?

Molar mass of empirical unit is 97.11 g/mol
Therefore the compound contains $\frac{190}{97.11} = 1.95 \approx 2$ empirical units
The compound is then $(C_4H_5N_2O)_2 = C_8H_{10}N_4O_2$

9) a) A lab student decomposed 1.241 g of $AlCl_3$ and was left with 0.262 g of Al. From the experimental data, what mass of Cl was in the original sample?

By conservation of mass, mass Cl in $AlCl_3$ = 1.241 g − 0.262 g = 0.979 g

b) According to the student's data, what is the experimental weight percent of Cl in $AlCl_3$?

$$\frac{0.979\text{ g}}{1.241\text{ g}} \times 100 = 78.9\ \%\text{ Cl by mass}$$

c) According to the chemical formula for $AlCl_3$, what is the actual (true) weight percent of Cl in $AlCl_3$?

$$\frac{3(35.45)}{26.98 + 3(35.45)} \times 100 = 79.76\ \%$$

d) To the correct significant figures, what is the percent error of the student's weight percent of Cl?

$$\frac{78.9 - 79.76}{79.76} \times 100 = \frac{-0.86}{79.76} \times 100 = -1\ \%$$

10) Balance the equation for the combustion of C_3H_7OH

$$2\ C_3H_7OH + 9\ O_2 \rightarrow 6\ CO_2 + 8\ H_2O$$

The next 5 questions refer to the following balanced equation:

$$Mg_3N_2(s) + 6\ H_2O(l) \rightarrow 3\ Mg(OH)_2(s) + 2\ NH_3(aq)$$

11) What mass of ammonia can be formed by reacting 10.0 g of Mg_3N_2?

$$\frac{10.0\text{ g } Mg_3N_2}{100.95\text{ g/mol}} = 0.0991\text{ mol } Mg_3N_2 \times \frac{2\text{ mol } NH_3}{1\text{ mol } Mg_3N_2} \times \frac{17.04\text{ g}}{\text{mol } NH_3} = 3.38\text{ g } NH_3$$

12) What mass of water is necessary to produce 4.5 x 10^{25} magnesium hydroxide formula units?

$$\frac{4.5 \times 10^{25}\ Mg(OH)_2\text{ units}}{6.022 \times 10^{23}\text{ units/mol}} = 75\text{ mol } Mg(OH)_2 \times \frac{6\text{ mol } H_2O}{3\text{ mol } Mg(OH)_2} \times \frac{18.02\text{ g}}{\text{mol } H_2O} = 2.7 \text{ x } 10^3\text{ g } H_2O$$

13) What mass of magnesium hydroxide can be formed by reacting 20.0 g of H_2O with 10.0 g of Mg_3N_2?

$$\frac{20.0\text{ g } H_2O}{18.02\text{ g/mol}} = 1.11\text{ mol } H_2O \qquad \frac{10.0\text{ g } Mg_3N_2}{100.95\text{ g/mol}} = 0.0991\text{ mol } Mg_3N_2$$

Check to see which is limiting: 1.11/ 6 = 0.185; 0.0991/1 = 0.0991 Since 0.0991 < 0.185, Mg_3N_2 limits

$$\frac{10.0\text{ g } Mg_3N_2}{100.95\text{ g/mol}} = 0.0991\text{ mol } Mg_3N_2 \times \frac{3\text{ mol } Mg(OH)_2}{1\text{ mol } Mg_3N_2} \times \frac{58.33\text{ g}}{\text{mol } Mg(OH)_2} = 17.3\text{ g}$$

14) What is the mass of excess reagent (i.e. reactant) leftover in Question 13?

$$\frac{10.0\text{ g } Mg_3N_2}{100.95\text{ g/mol}} = 0.0991\text{ mol } Mg_3N_2 \times \frac{6\text{ mol } H_2O}{1\text{ mol } Mg_3N_2} \times \frac{18.02\text{ g}}{\text{mol } H_2O} = 10.7\text{ g } H_2O \text{ used}$$

Therefore 20.0 g − 10.7 g = $\boxed{9.3\text{ g } H_2O \text{ left}}$

15) What is the percent yield of the reaction if only 10.0 g of magnesium hydroxide were collected from the reaction in Question 13?

$$\frac{10.0\text{ g}}{17.3\text{ g}} \times 100 = 57.8\ \%$$

16) The next 2 questions refer to the balanced reaction **$5\ O_2(g) + 2\ NiS_2(s) \rightarrow 2\ NiO(s) + 4\ SO_2(g)$**

a) Calculate the theoretical yield of SO_2 if 9.86 g of O_2 are reacted with 11.20 g of NiS_2.

Since you're given quantities of more than one reactant, you must see which one is limiting:

Step 1: Find moles of both: $\frac{9.86g}{32.00\text{g/mol}} = 0.308\text{ mol } O_2$ $\frac{11.20\text{g}}{122.83\text{g/mol}} = 0.0912\text{ mol } NiS_2$

Step 2: Divide each mole number by its stoichiometric coefficient. The smallest result is the limiting reactant

0.308/5 = 0.0616; 0.0912/2 = 0.0456

Since 0.0456 < 0.0616, NiS_2 is the limiting reactant.

Step 3: Use the moles of the limiting reactant to find the theoretical yield of product

$$\frac{11.20\text{g}}{122.83\text{g/mol}} = 0.0912\text{ mol } NiS_2 \times \frac{4\text{ mol } SO_2}{2\text{ mol } NiS_2} \times \frac{64.06\,g}{mol\ SO_2} = 11.68\text{ g } SO_2$$

b) If the actual yield of SO_2 is 10.00 g, what is percent yield of the reaction?

$$\frac{10.00}{11.68} \times 100 = 85.62\ \%$$