Stoichiometry Review - 11A

1. What volume of oxygen gas at STP can be prepared by the decomposition of 25.0 g of potassium chlorate? $2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$

$$\left(\frac{25.0\text{ g } KClO_3}{122.5495\text{ g/mol}}\right)\left(\frac{3\text{ mol } O_2}{2\text{ mol } KClO_3}\right)\left(22.4\text{ L/mol}\right) = 6.85\text{ L } O_2$$

2. Determine the volume of hydrogen gas at STP that can be produced by the reaction of 130.0 g of zinc with 100.0 g of hydrochloric acid. Which reactant is in excess, and how much remains unreacted? $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

$$\left(\frac{130.0\text{ g } Zn}{65.38\text{ g/mol}}\right)\left(\frac{1\text{ mol } H_2}{1\text{ mol } Zn}\right)\left(22.4\text{ L/mol}\right) = 44.5\text{ L } H_2$$

$$\left(\frac{100.0\text{ g } HCl}{36.46094\text{ g/mol}}\right)\left(\frac{1\text{ mol } H_2}{2\text{ mol } HCl}\right)\left(22.4\text{ L/mol}\right) = 30.7\text{ L } H_2$$

$\therefore$ 30.7 L H_2 produced

HCl is limiting, Zn is in excess

$$\left(\frac{100.0\text{ g } HCl}{36.46094\text{ g/mol}}\right)\left(\frac{1\text{ mol } Zn}{2\text{ mol } HCl}\right)\left(65.38\text{ g/mol}\right) = 89.7\text{ g } Zn \text{ needed}$$

$130.0\text{ g} - 89.7\text{ g} = 40.3\text{ g } Zn$ in excess

3. 29.6 mL of 0.350 M sodium hydroxide are required to titrate 20.0 mL of phosphoric acid. Calculate the molarity of the phosphoric acid. $3\ NaOH + H_3PO_4 \rightarrow Na_3PO_4 + 3H_2O$

$$C_a = \frac{C_b V_b R_a}{V_a R_b} = \frac{(0.350\text{ mol/L})(0.0296\text{ L})(1)}{(0.0200\text{ L})(3)}$$

$$= 0.173\text{ mol/L } H_3PO_4$$

4. Ammonium sulfate fertilizer is manufactured by having sulfuric acid react with ammonia. In a laboratory study of this process, 50.0 mL of sulfuric acid reacts with 24.4 mL of a 2.20 mol/L ammonium hydroxide solution. From this evidence, calculate the concentration of the sulfuric acid at this stage in the process.

$$H_2SO_4 + 2\,NH_4OH \rightarrow (NH_4)_2SO_4 + 2\,H_2O$$

$$C_a = \frac{C_b V_b R_a}{V_a R_b} = \frac{(2.20\ \text{mol/L})(0.0244\ \text{L})(1)}{(0.0500\ \text{L})(2)}$$

$$= 0.537\ \text{mol/L}\ H_2SO_4$$

5. Slaked lime can be added to an aluminum sulfate solution in a water treatment plant to clarify the water. Fine particles in the water stick to the precipitate produced. Calculate the volume of 0.0250 mol/L calcium hydroxide solution required to react completely with 25.0 mL of 0.125 mol/L aluminum sulfate solution.

$$3\,Ca(OH)_2 + Al_2(SO_4)_3 \rightarrow 2\,Al(OH)_3 + 3\,CaSO_4$$

$$V_a = \frac{C_b V_b R_a}{C_a R_b} = \frac{(0.125\ \text{mol/L})(0.0250\ \text{L})(3)}{(0.0250\ \text{mol/L})(1)}$$

$$= 0.375\ \text{L or } 375\ \text{mL}\ Ca(OH)_2$$

6. Some antacid products contain aluminum hydroxide to neutralize excess stomach acid. Determine the volume of 0.100 mol/L stomach acid (assumed to be HCl) that can be neutralized by 912 mg of aluminum hydroxide in an antacid tablet.

$$3\,HCl + Al(OH)_3 \rightarrow AlCl_3 + 3\,H_2O$$

$$\left(\frac{0.912\ \text{g}\ Al(OH)_3}{78.00356\ \text{g/mol}}\right)\left(\frac{3\ \text{mol HCl}}{1\ \text{mol}\ Al(OH)_3}\right) = 0.0351\ \text{mol HCl}$$

$$C = \frac{n}{V} \quad V = \frac{n}{C} = \frac{0.0351\ \text{mol}}{0.100\ \text{mol/L}} = 0.351\ \text{L or } 351\ \text{mL}$$

$Al(OH)_3$

7. **Iron(II) sulfide reacts with oxygen gas to produce iron(III) oxide and sulfur dioxide. What mass of iron(III) oxide is produced from the reaction of 20.0 g of iron(II) sulfide and 14.1 g of oxygen? Which reactant is in excess, and how much remains unreacted?**

$$4\,FeS + 7\,O_2 \rightarrow 2\,Fe_2O_3 + 4\,SO_2$$

$$\left(\frac{20.0\text{ g FeS}}{87.913\text{ g/mol}}\right)\left(\frac{2\text{ mol }Fe_2O_3}{4\text{ mol FeS}}\right)\left(159.6922\text{ g/mol}\right) = 18.2\text{ g }Fe_2O_3$$

$$\left(\frac{14.1\text{ g }O_2}{31.9988\text{ g/mol}}\right)\left(\frac{2\text{ mol }Fe_2O_3}{7\text{ mol }O_2}\right)\left(159.6922\text{ g/mol}\right) = 20.1\text{ g }Fe_2O_3$$

$\therefore$ 18.2 g Fe_2O_3 is produced

FeS is limiting, O_2 is in excess

$$\left(\frac{20.0\text{ g FeS}}{87.913\text{ g/mol}}\right)\left(\frac{7\text{ mol }O_2}{4\text{ mol FeS}}\right)\left(31.9988\text{ g/mol}\right) = 12.7\text{ g }O_2\text{ needed}$$

$$14.1\text{ g} - 12.7\text{ g} = 1.4\text{ g }O_2\text{ in excess}$$

8. **What is the maximum mass of sodium chloride that can be produced when 50.0 mL of 0.120 mol/L sodium hydroxide reacts with 39.4 mL of 0.165 mol/L hydrochloric acid?**

$$NaOH + HCl \rightarrow NaCl + H_2O$$

$$n = C\cdot V = (0.0500\text{ L})(0.120\text{ mol/L}) = 0.00600\text{ mol NaOH}$$

$$(0.00600\text{ mol NaOH})\left(\frac{1\text{ mol NaCl}}{1\text{ mol NaOH}}\right)\left(58.44277\text{ g/mol}\right) = 0.351\text{ g NaCl}$$

$$n = C\cdot V = (0.0394\text{ L})(0.165\text{ mol/L}) = 0.006501\text{ mol HCl}$$

$$(0.006501\text{ mol HCl})\left(\frac{1\text{ mol NaCl}}{1\text{ mol HCl}}\right)\left(58.44277\text{ g/mol}\right) = 0.380\text{ g NaCl}$$

$\therefore$ 0.351 g NaCl is produced

NaOH is limiting, HCl is in excess

9. A chemical technician analyzed a sample of a waste solution and reported findings to environmental chemists monitoring the lead content of the waste. The sample was mixed with excess potassium bromide so that any lead(II) nitrate in the solution would react; 3.65 g of precipitate was produced. Calculate the mass of lead(II) nitrate in the sample. $Pb(NO_3)_2 + 2KBr \rightarrow 2KNO_3 + PbBr_2 \downarrow$

$$\left(\frac{3.65\,g\ PbBr_2}{367.008\ g/mol}\right)\left(\frac{1\ mol\ Pb(NO_3)_2}{1\ mol\ PbBr_2}\right)\left(331.2098\ g/mol\right)$$

$$= 3.29\ g\ Pb(NO_3)_2$$

10. A chromium(III) chloride solution is analyzed by having a sample of the solution react with a 50.0 g piece of zinc metal. After the reaction, 38.5 g of zinc remained. Calculate the mass of chromium(III) chloride that was present in the sample tested.

$$2\,CrCl_3 + 3\,Zn \rightarrow 3\,ZnCl_2 + 2\,Cr$$

50.0 g initial
− 38.5 g remains
11.5 g Zn used

$$\left(\frac{11.5\,g\ Zn}{65.38\ g/mol}\right)\left(\frac{2\ mol\ CrCl_3}{3\ mol\ Zn}\right)\left(158.355\ g/mol\right)$$

$$= 18.6\ g\ CrCl_3$$

11. Sulfuric acid is produced on a large scale from readily available raw materials. One step in the industrial production of sulfuric acid is the reaction of sulfur trioxide with water. Calculate the molar concentration of sulfuric acid produced by the reaction of 10.0 Mg of sulfur trioxide with an excess quantity of water to produce 7.00 kL of acid.

$$SO_3 + H_2O \rightarrow H_2SO_4$$

$10.0\ Mg \left(\frac{10^6\ g}{1\ Mg}\right) = 1.00 \times 10^7\ g\ SO_3$ $7.00\ kL\left(\frac{1000\ L}{1\ kL}\right) = 7.00 \times 10^3\ L$ acid

$\left(\frac{1.00 \times 10^7\ g\ SO_3}{80.0642\ g/mol}\right)\left(\frac{1\ mol\ H_2SO_4}{1\ mol\ SO_3}\right) = 124\,899.8\ mol\ H_2SO_4$ $C = \frac{n}{V} = \frac{124\,899.8\ mol}{7.00 \times 10^3\ L}$

$= 17.8\ mol/L$

12. A 1.250 g sample of the compound responsible for the odour of cloves, which is known to contain only the elements C, H, and O, is burned in oxygen to produce carbon dioxide and water. The mass of carbon dioxide produced is 3.350 g, and the mass of water produced is 0.8232 g. Determine the empirical and molecular formulas of this compound, given that its molecular mass is 164 g/mol.

$$C_?H_?O_? + O_2 \rightarrow CO_2 + H_2O$$

CO_2 $\frac{3.350\ g}{44.0098\ g/mol}$

$= 0.0761\ mol\ CO_2$

$= 0.0761\ mol\ C$

$\times 12.011\ g/mol$

$= 0.914\ g\ C$

1.250 g total
−0.914 g C
−0.0921 g H

0.244 g O

H_2O $\frac{0.8232\ g}{18.0152 8\ g/mol}$

$= 0.0457\ mol\ H_2O$

$\times 2$

$= 0.0914\ mol\ H$

$\times 1.00794$

$= 0.0921\ g\ H$

$\frac{0.914\ g\ C}{12.011\ g/mol}$ $\frac{0.0921\ g\ H}{1.00794\ g/mol}$ $\frac{0.244\ g\ O}{15.9994\ g/mol}$

$\frac{0.0761\ mol}{0.0152\ mol}$ $\frac{0.0914\ mol}{0.0152\ mol}$ $\frac{0.0152\ mol}{0.0152\ mol}$

5 : 6 : 1

C_5H_6O empirical $\frac{\text{mol mass}}{\text{emp mass}} = \frac{164}{82.102} = 2$

$\therefore C_{10}H_{12}O_2$ molecular

13. A 0.473 g sample of phosphorus is reacted with an excess of chlorine, and 2.12 g of phosphorus pentachloride is collected. What is the percentage yield of phosphorus pentachloride? $2P + 5Cl_2 \rightarrow 2PCl_5$

$$\left(\frac{0.473\text{ g P}}{30.973\,76\text{ g/mol}}\right)\left(\frac{2\text{ mol }PCl_5}{2\text{ mol P}}\right)\left(208.238\,76\text{ g/mol}\right) = 3.18\text{ g}$$

$$\%\text{ yield} = \frac{2.12\text{ g}}{3.18\text{ g}} \times 100 = 66.7\%\text{ yield }PCl_5$$

14. Butylated hydroxytoluene, BHT, a food preservative, contains carbon, hydrogen, and oxygen. A 15.42 mg sample of BHT was burned in a stream of oxygen and yielded 46.20 mg of carbon dioxide and 15.13 mg of water. Calculate the empirical formula of BHT. $C_?H_?O_? + O_2 \rightarrow CO_2 + H_2O$

CO_2 $\dfrac{0.04620\text{ g}}{44.0098\text{ g/mol}}$

$= 0.00105\text{ mol }CO_2$

$= 0.00105\text{ mol C}$

$\times 12.011\text{ g/mol}$

$= 0.01261\text{ g C}$

H_2O $\dfrac{0.01513\text{ g}}{18.01528\text{ g/mol}}$

$= 0.000\,8398\text{ mol }H_2O$

$\times 2$

$= 0.00168\text{ mol H}$

$\times 1.00794\text{ g/mol}$

$= 0.00169\text{ g H}$

$\underset{\text{total}}{0.01542\text{ g}} - \underset{\text{C}}{0.01261\text{ g}} - \underset{\text{H}}{0.00169\text{ g}}$

$= 0.00112\text{ g O}$

$\dfrac{0.01261\text{ g C}}{12.011\text{ g/mol}}$ $\dfrac{0.00169\text{ g H}}{1.00794\text{ g/mol}}$ $\dfrac{0.00112\text{ g O}}{15.9994\text{ g/mol}}$

$\dfrac{0.00105\text{ mol}}{6.99\times10^{-5}}$ $\dfrac{0.00168\text{ mol}}{6.99\times10^{-5}}$ $\dfrac{6.99\times10^{-5}\text{ mol}}{6.99\times10^{-5}}$

15 : 24 : 1

$C_{15}H_{24}O$

15. Pyridine is recovered from coke-oven gases and is used extensively in the chemical industry, in particular, in the synthesis of vitamins and drugs. Pyridine contains carbon, hydrogen, and nitrogen. A 0.546 g sample was burned to produce 1.518 g of carbon dioxide and 0.311 g of water. Determine the empirical formula of pyridine.

$$C_? H_? N_? + O_2 \rightarrow CO_2 + H_2O + ?$$

0.546 g — 1.518 g — 0.311 g

CO_2 $\frac{1.518 \text{ g}}{44.0098 \text{ g/mol}}$

$= 0.03449 \text{ mol } CO_2$

$= 0.03449 \text{ mol C}$

$\times 12.011 \text{ g/mol}$

$= 0.414 \text{ g C}$

H_2O $\frac{0.311 \text{ g}}{18.01528 \text{ g/mol}}$

$= 0.01726 \text{ mol } H_2O$

$\times 2$

0.03453 mol H

$\times 1.00794$

$= 0.0348 \text{ g H}$

0.546 g total
- 0.414 g C
- 0.0348 g H

0.0969 g N

$\frac{0.414 \text{ g C}}{12.011 \text{ g/mol}}$	$\frac{0.0348 \text{ g H}}{1.00794 \text{ g/mol}}$	$\frac{0.0969 \text{ g N}}{14.0067 \text{ g/mol}}$
$\frac{0.03449 \text{ mol}}{0.00691}$	$\frac{0.03453 \text{ mol}}{0.00691}$	$\frac{0.00691 \text{ mol}}{0.00691}$
5 :	5 :	1

C_5H_5N

16. Nitric acid is made commercially from ammonia by the Ostwald process, which was developed by the German chemist Wilhelm Ostwald. The process consists of three steps. In the first step ammonia reacts with oxygen to produce nitrogen monoxide and water. In the second step nitrogen monoxide is reacted with oxygen and converted to nitrogen dioxide. In the third and final step, the nitrogen dioxide is reacted with water to produce nitric acid and nitrogen monoxide. What mass of nitric acid can be produced from 6.40 x 10^4 kg of ammonia?

$$4\ NH_3 + 5\ O_2 \rightarrow 4\ NO + 6\ H_2O$$

$$2\ NO + O_2 \rightarrow 2\ NO_2$$

$$3\ NO_2 + H_2O \rightarrow 2\ HNO_3 + NO$$

$$\left(\frac{6.40 \times 10^4\ kg\ NH_3}{17.03052\ kg/kmol}\right)\left(\frac{4\ mol\ NO}{4\ mol\ NH_3}\right) = 3757.96\ kmol\ NO$$

$$(3757.96\ kmol\ NO)\left(\frac{2\ mol\ NO_2}{2\ mol\ NO}\right) = 3757.96\ kmol\ NO_2$$

$$(3757.96\ kmol\ NO_2)\left(\frac{2\ mol\ HNO_3}{3\ mol\ NO_2}\right)\left(63.01284\ kg/kmol\right)$$

$$= 157\,866.4562\ kg$$

$$= 1.58 \times 10^5\ kg\ HNO_3$$

(or 158 Mg)

17. The combustion of gasoline in an automobile engine can be represented by the equation:

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

a) In a properly tuned engine with a full tank of gas, what reactant do you think is limiting? Explain your reasoning.

The gasoline would be the limiting reactant. Gasoline is a non-renewable fuel and must be purchased. Oxygen is free and readily available in the air surrounding the car. It is much more practical to design the system so that the gasoline is the limiting reactant.

b) A car that is set to inject the correct amount of fuel at sea level will run poorly at higher altitudes, where the air is less dense. Explain why.

At higher altitudes, air is less dense, therefore the amount of oxygen is the same volume of air will be less than it would be at sea level. This means that the fuel to air ratio will not be at optimum levels at the higher altitude.

c) Calculate the volume of carbon dioxide gas (at STP) that is produced by the combustion of 10.0 g of gasoline.

$$\left(\frac{10.0\text{ g } C_8H_{18}}{114.23092\text{ g/mol}}\right)\left(\frac{16\text{ mol } CO_2}{2\text{ mol } C_8H_{18}}\right)\left(22.4\text{ L/mol}\right)$$

$$= 15.6875$$

$$= 15.7\text{ L } CO_2$$

c) The reaction of atmospheric oxygen with atmospheric nitrogen to form nitrogen monoxide occurs in a car's engine along with the combustion of fuel.

$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$

What adjustments need to be made to a vehicle's fuel injectors (which control the amount of fuel and air that are mixed) to compensate for this reaction? Explain your answer.

The combustion of nitrogen uses up some of the available oxygen, this means that there is less oxygen available for the combustion of the gasoline. Because of this, the amount of air that enters the engine must be increased, so that sufficient amount of oxygen is available for the combustion of the gasoline.

18. An impure sample of barium chloride with a mass of 4.36 g is added to an aqueous solution of sodium sulfate.

$$BaCl_2(s) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$

After the reaction is complete, the solid barium sulfate produced by the reaction is filtered and dried. Its mass is found to be 2.62 g. What is the percentage purity of the original barium chloride?

$$\left(\frac{2.62\text{g } BaSO_4}{233.3936 \text{ g/mol}}\right)\left(\frac{1\text{ mol } BaCl_2}{1\text{ mol } BaSO_4}\right)\left(208.236\frac{\text{g}}{\text{mol}}\right)$$

$$= 2.34 \text{ g } BaCl_2$$

$$\% \text{ purity} = \frac{2.34\text{g}}{4.36\text{g}} \times 100 = 53.6\% \text{ pure}$$

19. Complex carbohydrates are starches that your body can convert to glucose, a type of sugar. Simple carbohydrate foods contain glucose, ready for immediate use by the human body. Breathing and burning glucose produces energy in a jogger's muscles, according to the following unbalanced equation:

$$C_6H_{12}O_6(aq) + O_2(g) \rightarrow CO_2(g) + H_2O(g) + \text{energy}$$

Just before going on a run, Marina eats tow oranges. The oranges give her body 25 g of glucose to make energy. Although 21.0% (by volume) of the air Marina breathes in is oxygen, she breathes out about 16.0% of this oxygen. (In other words, she only uses about 5.0%.)

a) Balance the equation

$$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g) + \text{energy}$$

b) How many litres of air (at STP) does Marina breathe in while running to burn up the glucose she consumed?

$$\left(\frac{25\text{ g } C_6H_{12}O_6}{180.15768 \text{ g/mol}}\right)\left(\frac{6\text{ mol } O_2}{1\text{ mol } C_6H_{12}O_6}\right)\left(22.4\frac{\text{L}}{\text{mol}}\right) = 18.6503\text{L } O_2 \text{ needed}$$

she uses 5% of the air (source of O_2)

$$\frac{5.0}{100.0} = \frac{18.650}{x} \qquad x = 373.0066$$

∴ she needs 370 L of air

c) How many litres of carbon dioxide (at STP) does she produce?

$$\left(\frac{25.0\text{g } C_6H_{12}O_6}{180.15768 \text{ g/mol}}\right)\left(\frac{6\text{ mol } CO_2}{1\text{ mol } C_6H_{12}O_6}\right)\left(22.4\frac{\text{L}}{\text{mol}}\right)$$

$$= 18.7 \text{ L } CO_2$$