General Chemistry I
PLI Sheet 9 AK
June 15, 2021

1. **Nitric oxide is made from the oxidation of ammonia. What mass of nitric oxide can be made from the reaction of 8.00 g NH_3 with 17.0 g O_2 given the following, unbalanced equation? If 10. grams of NO are obtained, what is the percent yield for the reaction?**

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

$$8.00\ g\ NH_3 * \frac{1\ mol\ NH_3}{17.03\ g\ NH_3} * \frac{4\ mol\ NO}{4\ mol\ NH_3} = 0.4698\ mol\ NO$$

$$17.0\ g\ O_2 * \frac{1\ mol\ O_2}{31.9988\ g\ O_2} * \frac{4\ mol\ NO}{5\ mol\ O_2} = 0.425\ mol\ NO$$

O_2 is the limiting reactant since it produces less NO. Calculate the mass of NO that can be theoretically obtained:

$$0.425\ mol\ NO * \frac{30.01\ g\ NO}{1\ mol\ NO} = 12.75\ g\ NO$$

$$\%\ yield = \frac{10.\ g\ NO}{12.75\ g\ NO} * 100\% = 78\%\ yield$$

2. **You mix aqueous silver nitrate with aqueous aluminum chloride and a precipitate forms! You use 200.0 g of aluminum chloride and 325 g of silver nitrate.**
   a. **Write the balanced and net ionic equations.**

$$3AgNO_3\,(aq) + AlCl_3\,(aq) \rightarrow 3AgCl\,(s) + Al(NO_3)_3(aq)$$

$$3Ag^+\,(aq) + 3NO_3^-\,(aq) + Al^{3+}\,(aq) + 3Cl^-\,(aq) \rightarrow 3AgCl\,(s) + Al^{3+}\,(aq) + 3NO_3^-(aq)$$

$$3Ag^+\,(aq) + 3Cl^-\,(aq) \rightarrow 3AgCl\,(s)$$

   b. **How much solid can theoretically be produced?**

This is a limiting reactant problem!

$$200.0\ g\ AlCl_3 * \frac{1\ mol\ AlCl_3}{133.34\ g\ AlCl_3} * \frac{3\ mol\ AgCl}{1\ mol\ AlCl_3} = 4.50\ mol\ AgCl$$

$$325\ g\ AgNO_3 * \frac{1\ mol\ AgNO_3}{169.87\ g\ AgNO_3} * \frac{3\ mol\ AgCl}{3\ mol\ AgNO_3} = 1.91\ mol\ AgCl$$

$AgNO_3$ is the limiting reactant.

$$1.91\ mol\ AgCl * \frac{143.32\ g\ AgCl}{1\ mol\ AgCl} = 274.2\ g\ AgCl$$

c. **After doing the experiment, you got a 92.0% yield. How many chlorine atoms are there in the solid that you made?**

$$274.2\ g\ AgCl * 0.92 = 252.27\ g\ AgCl\ formed$$

$$252.27\ g\ AgCl * \frac{1\ mol\ AgCl}{143.32\ g\ AgCl} * \frac{1\ mol\ Cl^-}{1\ mol\ AgCl} * \frac{6.02*10^{23}\ atoms\ Cl^-}{1\ mol\ Cl^-} = 1.06 * 10^{24}\ Cl^-\ atoms$$

3. **The commercial production of nitric acid involves the following reactions. Balance them and identify which ones are redox reactions. For the redox reactions, identify each element's oxidation number, the oxidizing agent (OA), and the reducing agent (RA)**

$$4NH_3\,(g) + 5O_2\,(g) \rightarrow 4NO\,(g) + 6H_2O(g)$$

N: -3 to +2 (oxidized); O: 0 to -2 (reduced); OA: O_2; RA: NH_3

$$2NO\,(g) + O_2\,(g) \rightarrow 2NO_2(g)$$

N: +2 to +4 (oxidized); O: 0 to -2 (reduced); OA: O_2; RA: NO

$$3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO(g)$$

N: +4 to +5 (oxidized); N: +4 to +2 (reduced); OA & RA are NO_2

All three reactions are redox reactions!

4. **An oxoacid with the formula $H_xE_yO_z$ has a formula mass of 178 g/mol, has 13 atoms in its formula unit, and contains 34.80% by mass, and 15.38% by number of atoms, of the element E. What is the element E, and what is the formula of the oxoacid?**

$$13\ atoms * 0.1538 = 2\ atoms\ of\ E\ in\ formula\ unit$$

$$Mass\ E\ in\ formula = 0.3480 * \frac{178\ g}{mol} = 61.94 \frac{g}{mol}$$

We can find the molar mass of E by dividing the mass of E in the formula by 2 since there are 2 atoms of E.

$$\frac{61.94\frac{g}{mol}}{2} = 30.97\frac{g}{mol} \rightarrow E \text{ must be Phosphorus}$$

The rest of the molar mass of the compound must be for the H and O in the formula.

$$Mass\ of\ H\ \&\ O = 178\frac{g}{mol} - 61.94\frac{g}{mol} = 116.06\frac{g}{mol}$$

Remember that x and z represent the number of atoms of H and O in the formula, respectively.

$$\#\ atoms\ H + \#\ atoms\ O = 13 - 2 = 11$$

$$x + z = 11$$

$$x\left(1.008\frac{g}{mol}\right) + z\left(15.9994\frac{g}{mol}\right) = 116.06\frac{g}{mol}$$

$$x(1.008) + (11 - x)(15.9994) = 116.06$$

$$1.008x + 175.9934 - 15.9994x = 116.06$$

$$-14.9914x = -59.9334$$

$$x = 4\ so\ z = 7$$

Formula: $H_4P_2O_7$

5. **Find ΔH_{rxn} for the reaction: $3C(s) + 4H_2(g)$ □ $C_3H_8(g)$. Use these reactions with known ΔH.**

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g) \qquad \Delta H = -2043\ kJ$$

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -393.5\ kJ$$

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g) \qquad \Delta H = -483.6\ kJ$$

We want C_3H_8 on the products side, but in the first equation it is on the reactant side. So, reverse the first equation (like multiplying by -1).

$$3CO_2(g) + 4H_2O(g) \rightarrow C_3H_8(g) + 5O_2(g) \qquad \Delta H = 2043\ kJ$$

Next, we want 3 moles of C(s). So multiply the second equation by 3 to become:

$$3C\,(s) + 3O_2\,(g) \rightarrow 3CO_2\,(g) \qquad \Delta H = -\,1180.5\ kJ$$

Finally, we want 4 moles of H_2(g) on the reactant side. Multiply the third equation by 2.

$$4H_2\,(g) + 2O_2\,(g) \rightarrow 4H_2O\,(g) \qquad \Delta H = -\,967.2\ kJ$$

Finally, sum all three equations together and the corresponding enthalpy values to get:

$$3C\,(s) + 4H_2\,(g) \rightarrow C_3H_8\,(g) \qquad \Delta H = -\,104.7\ kJ$$

6. **When 5.03 g of solid potassium hydroxide are dissolved in 100.0 mL of distilled water in a coffee-cup calorimeter, the temperature of the liquid increases from 23.0°C to 34.7°C (density of water 1.00 g/cm^3). What is the solution ΔH (the enthalpy for dissolving) in kJ per mole of KOH? Assume the calorimeter absorbs a negligible amount of heat, and because of the large volume of water, that the specific heat of the solution is the same as that of pure water (4.184 J/ g °C).**

   The premise here is similar to what was covered in an earlier lecture when thinking about dropping a hot metal in water and considering the heat transfer. The heat from dissolving the potassium hydroxide is transferred directly to the solution.

$$q_{dissolving} + q_{solution} = 0$$

$$q_{solution} = -\,q_{dissolving}$$

To find $q_{dissolving}$, first find the mass of the solution:

$$Total\ mass = \left(100.0\ mL * 1.00\tfrac{g}{mL}\right) + 5.03\ g = 105.03\ g$$

$$q_{dissolving} = mc\Delta T = (105.03\ g)\left(4.184\tfrac{J}{g℃}\right)(34.7℃ - 23.0℃) = 5141.5\ J = 5.14\ kJ$$

Thus, $q_{solution}$ is equal to = -5.14 kJ. Since the problem asks for kJ per mole of KOH, we need to find the moles of KOH.

$$5.03\ g\ KOH * \frac{1\ mol\ KOH}{56.11\ g\ KOH} = 0.0896\ mol\ KOH$$

$$q_{solution} = \frac{-5.14\ kJ}{0.0896\ mol\ KOH} = -\,57.4\,\frac{kJ}{mol\ KOH}$$

7. **Consider the following reaction:**

$$2Mg\,(s) + O_2\,(g) \rightarrow 2MgO\,(s) \qquad \Delta H = -1204\ kJ$$

a. **Is this reaction exothermic or endothermic?**
Exothermic! Enthalpy is negative

b. **Calculate the amount of heat transferred when 3.55 g of Mg(s) reacts at constant pressure.**

$$3.55\ g\ Mg * \frac{1\ mol\ Mg}{24.31\ g\ Mg} * \frac{1\ molrxn}{2\ mol\ Mg} * \frac{-1204\ kJ}{1\ molrxn} = -87.9\ kJ\ transferred$$

c. **How many grams of MgO are produced during an enthalpy change of -234 kJ?**

$$-234\ kJ * \frac{1\ molrxn}{-1204\ kJ} * \frac{2\ mol\ MgO}{1\ molrxn} * \frac{40.30\ g\ MgO}{1\ mol\ MgO} = 15.7\ g\ MgO$$

d. **How many kilojoules of heat are absorbed when 40.3 g of MgO(s) is decomposed into Mg(s) and O_2(g) at constant pressure?**

This is the reverse reaction! Make sure you flip the sign of the enthalpy

$$40.3\ g\ MgO * \frac{1\ mol\ MgO}{40.30\ g\ MgO} * \frac{1\ molrxn}{2\ mol\ MgO} * \frac{1204\ kJ}{1\ molrxn} = 602\ kJ\ absorbed$$

It should make sense that this is a positive number because heat is now being absorbed (endothermic reaction).

8. **Ethanol (C_2H_5OH) is currently blended with gasoline as an automobile fuel.**

a. **Write a balanced equation for the combustion of liquid ethanol in air.**

$$C_2H_5OH\,(l) + 3O_2\,(g) \rightarrow 2CO_2\,(g) + 3H_2O(g)$$

b. **Calculate the standard enthalpy change for the reaction, assuming H_2O(g) as a product. Some useful $\Delta_f H°$ values: H_2O(g): -241.82 kJ/mol; CO_2(g): -393.5 kJ/mol; C_2H_5OH(l): -277.7 kJ/mol.**

$$\Delta H°_{rxn} = \Sigma n \Delta H_f° \,(products) - \Sigma m \Delta H_f° \,(reactants)$$

$$\Delta H°_{rxn} = \left[\left(-241.82 \frac{kJ}{mol\ H_2O} * 3\ mol\ H_2O\right) + \left(-393.5 \frac{kJ}{mol\ CO_2} * 2\ mol\ CO_2\right)\right] - \left[0 + \left(-277.7 \frac{kJ}{mol\ C_2H_5OH} * 1\ m\right.\right.$$

*** 1 mol C2H5OH)] =**

-1234.8 kJ/mol rxn

c. **Calculate the heat produced per liter of ethanol by combustion of ethanol under constant pressure. Ethanol has a density of 0.789 g/mL.**

This is another way of asking how much heat is produced if 1 L of ethanol is combusted. First how many moles of ethanol is in 1 L of ethanol.

$$1\ L\ C_2H_5OH * \frac{1000\ mL}{1\ L} * 0.789\frac{g}{mL} * \frac{1\ mol\ C_2H_5OH}{46.07\ g\ C_2H_5OH} = 17.126\ mol\ C_2H_5OH$$

Then, use stoichiometry to find the proportional amount of heat produced.

$$17.126\ mol\ C_2H_5OH * \frac{1\ molrxn}{1\ mol\ C_2H_5OH} * \frac{-1234.8\ kJ}{1\ molrxn} =- 21147.3\ kJ =- 2.11 * 10^4\ kJ$$

Thus, heat produced per liter is $2.11*10^4$ kJ/L ethanol.

d. **Calculate the mass of CO_2 produced per kJ of heat emitted.**

Once again, this question is another way of asking how much CO_2 is produced if 1 kJ of heat is emitted. So, start off with 1 kJ of heat being emitted.

$$-1\ kJ * \frac{1\ molrxn}{-1234.8\ kJ} * \frac{2\ mol\ CO_2}{1\ molrxn} * \frac{44.01\ g\ CO_2}{1\ mol\ CO_2} = 0.0713\ g\ CO_2$$

Thus, the answer is 0.0713 g of CO_2/kJ heat emitted.

9.

a. **A strontium hydroxide solution is prepared by dissolving 12.50 g of $Sr(OH)_2$ in water to make 50.00 mL of solution. What is the molarity of this solution?**

$$12.50\ g\ Sr(OH)_2 * \frac{1\ mol\ Sr(OH)_2}{121.63\ g\ Sr(OH)_2} = 0.1028\ mol\ Sr(OH)_2$$

$$[Sr(OH)_2] = \frac{0.1028\ mol\ Sr(OH)_2}{0.05000\ L} = 2.055\ M$$

b. **Next the strontium hydroxide solution prepared in part (a) is used to titrate a nitric acid solution of unknown concentration. Write a balanced chemical equation to represent this reaction.**

$$Sr(OH)_2(aq) + 2HNO_3(aq) \rightarrow Sr(NO_3)_2(aq) + 2H_2O(l)$$

c. **If 23.9 mL of the strontium hydroxide solution was needed to neutralize a 37.5 mL sample of the nitric acid solution, what is the concentration of the acid?**

First, find out how many moles of base were added to neutralize the acid.

$$2.055\ M\ Sr(OH)_2 : \frac{2.055\ mol\ Sr(OH)_2}{1\ L} * 0.0239\ L = 0.0491\ mol\ Sr(OH)_2$$

Next, it is important to determine how much hydroxide (OH^-) was added, because we know the amount of acid must be equal to the amount of OH^- added for the full neutralization to occur.

$$0.0491\ mol\ Sr(OH)_2 * \frac{2\ mol\ OH^-}{1\ mol\ Sr(OH)_2} = 0.0982\ mol\ OH^-\ added$$

Thus, 0.0982 mol of acid were neutralized. Since HNO_3 just has one acidic proton, 0.0982 mol of HNO_3 were neutralized. Now, the HNO_3 concentration can be calculated.

$$[HNO_3] = \frac{0.0982\ mol\ HNO_3}{0.0375\ L} = 2.62\ M$$