PLI Worksheet #6
June 11, 2021

1. It takes 83 mL of a 0.45 M NaOH solution to neutralize 235 mL of an HCl solution. What is the concentration of the HCl solution?

Always write the balanced equation first!

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

$$\frac{0.45\ mol\ NaOH}{1\ L} * 0.083\ L = 0.03735\ mol\ NaOH * \frac{1\ mol\ OH^-}{1\ mol\ NaOH} * \frac{1\ mol\ HCl}{1\ mol\ OH^-} = 0.03735\ mol\ HCl$$

0.03735 mol of HCl was neutralized.

$$[HCl] = \frac{0.037535\ mol\ HCl}{0.235\ L} = 0.16\ M$$

2. A 0.5895-g sample of impure magnesium hydroxide is dissolved in 100.0 mL of 0.2050M HCl solution. The excess acid then needs 19.85 mL of 0.1020 M NaOH for neutralization. Calculate the percentage by mass of magnesium hydroxide in the sample, assuming that it is the only substance reacting with the HCl solution.

First, come up with the balanced equation: $2HCl+Mg(OH)_2 \rightarrow MgCl_2+2H_2O$

Then, we need to determine how many moles of HCl were neutralized by the initial dissolution of $Mg(OH)_2$. We can first determine how many moles of HCl were present in the solution, since we know its volume and molarity:

We also know how much leftover we had to neutralize using NaOH:

$$\frac{0.2050\ mol\ HCl}{1\ L} * 0.100\ L = 0.0205\ mol\ HCl\ present\ (total)$$

$$\frac{0.1020\ mol\ NaOH}{1\ L} * 0.01985\ L = 0.0020247\ mol\ NaOH\ to\ neutralize\ excess\ acid$$

So, the initial amount neutralized is:

Initial amount= (Total moles HCl)-(Excess Moles HCL neutralized using NaOH)

$0.0205\ mol\ HCl\ \text{total} - 0.0020247\ mol\ HCl\ excess = 0.0184753\ mol\ HCl\ initially\ neutralized$

Since we had excess acid to be neutralized, we know that we would need an equal amount of OH^- ions to accomplish this. So, we can determine the total mass of $Mg(OH)_2$ present using this information, and the formula for magnesium hydroxide. Then, we simply divide by the total mass of what was dissolved to determine the mass percent.

$$0.0184753\ mol\ OH^- * \frac{1\ mol\ Mg(OH)_2}{2\ mol\ OH^-} * \frac{58.32\ g\ Mg(OH)_2}{1\ mol\ Mg(OH)_2} = 0.5387\ g\ Mg(OH)_2$$

$$Mass\ \%\ Mg(OH)_2 = \frac{Mg(OH)_2\ mass}{Total\ mass} = \frac{0.5387\ g}{0.5895\ g} * 100\% = 91.39\%$$

3. An unknown solid acid is either citric acid or tartaric acid. To determine which acid you have, you titrate a sample of the solid with aqueous NaOH and from this determine the molar mass of the unknown acid. A 0.857g sample requires 26.1 mL of 0.513 M NaOH to consume the acid completely. What is the unknown acid? The appropriate equations are as follows:

Citric acid: $H_3C_6H_5O_{7(aq)} + 3NaOH_{(aq)} \longrightarrow 3H2O_{(l)} + Na_3C_6H_5O_{7(aq)}$

Tartaric acid: $H_2C_4H_4O_{(aq)} + 2\ NaOH_{(aq)} \longrightarrow 2H_2O_{(l)} + Na_2C_4H_4O_{6\ (aq)}$

First, determine the moles of NaOH used in this titration:

$$0.0261\ L\ NaOH * \left(\frac{0.513\ mol\ NaOH}{1L}\right) = 0.0134 \text{ mol NaOH used}$$

Then, check each equation to see in which this amount of NaOH would completely react with a mass of acid equal to .857g. In this case:

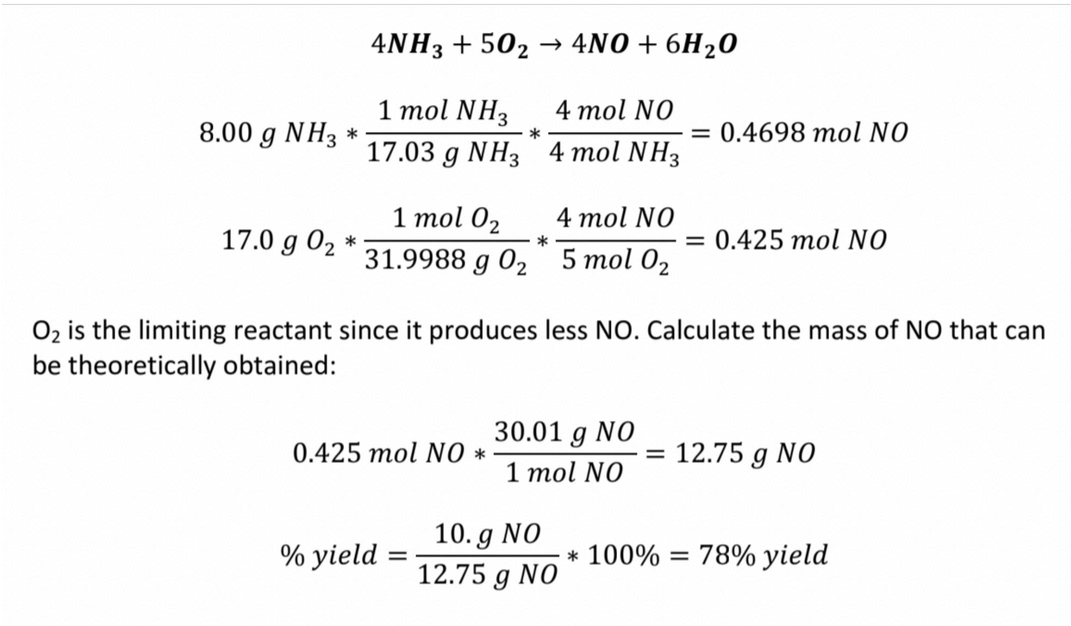
$$0.0134\ mol\ NaOH * \left(\frac{1\ mol\ citric\ acid}{3\ mol\ NaOH}\right) * \frac{192.124\ g\ citric\ acid}{1\ mol\ citric\ acid} = 0.857\ g\ citric\ acid$$

So, we can assume it’s citric acid. We can double check with the tartaric acid.

$$0.0134\ mol\ NaOH * \frac{1\ mol\ tartaric\ acid}{2\ mol\ NaOH} * \frac{70.09\ g\ tartaric\ acid}{1\ mol\ tartaric\ acid} = 0.470 \text{ g tartaric acid}$$

4. Nitric oxide is made from the oxidation of ammonia. What mass of nitric oxide can be made from the reaction of 8.00 g NH3 with 17.0 g O2 given the following, unbalanced equation? If 10.g of NO are obtained, what is the percent yield for the reaction?

___NH3 + ___O2 ⟶ ___NO + ___H2O



$$4\boldsymbol{NH_3} + 5\boldsymbol{O_2} \rightarrow 4\boldsymbol{NO} + 6\boldsymbol{H_2O}$$

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

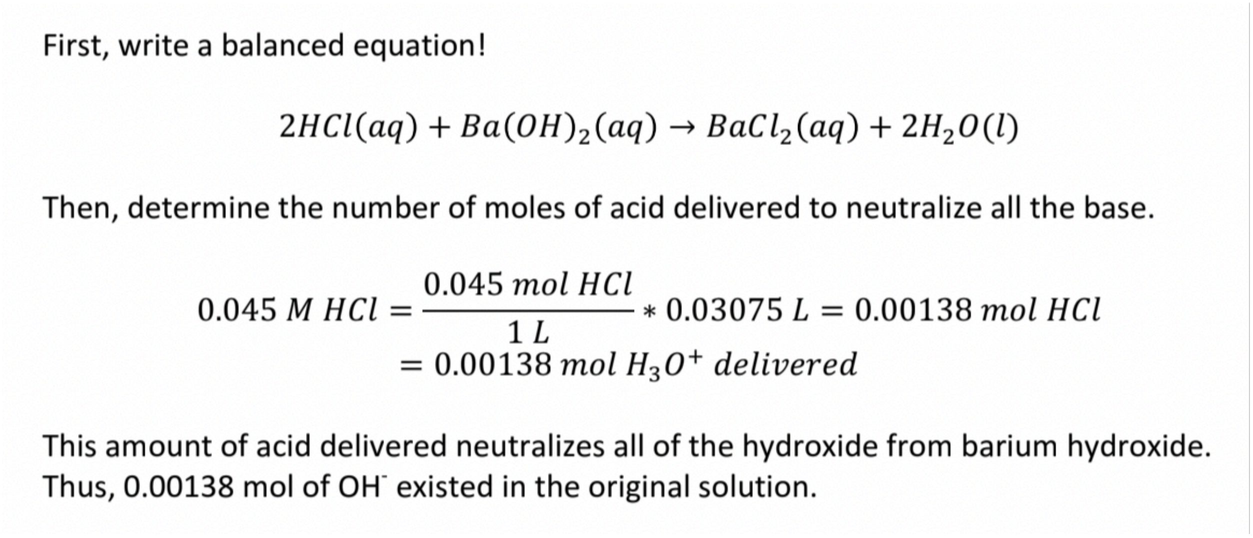
$$17.0\ g\ O_2 * \frac{1\ mol\ O_2}{31.9988\ g\ O_2} * \frac{4\ mol\ NO}{5\ mol\ O_2} = 0.425\ mol\ NO$$

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

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

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

5. You are given 100.0 mL of barium hydroxide solution that is of unknown concentration. You perform a titration using 0.045 M HCl and find that you reach the equivalence point of the titration after 30.75 mL of the acid was added. What was the molarity of the original barium hydroxide solution? What was the pH of the original solution?



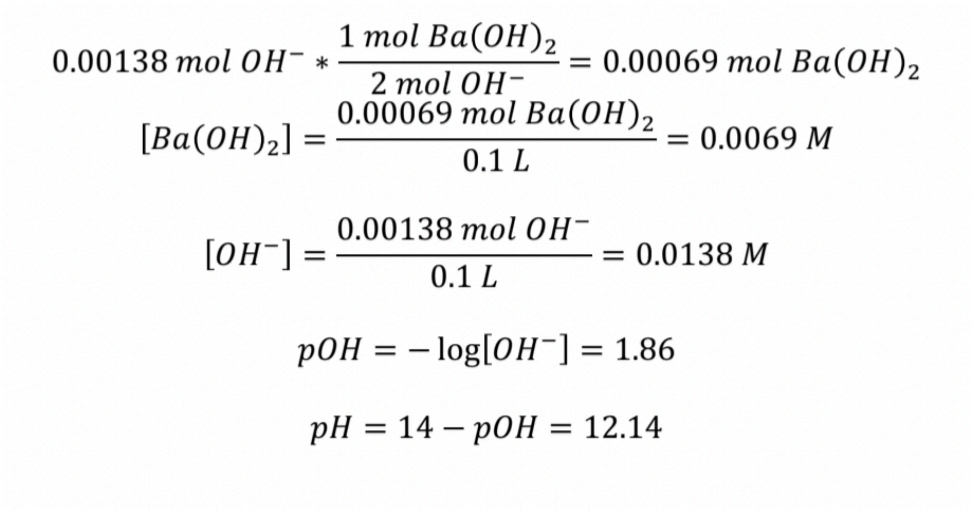
First, write a balanced equation!

$$2HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$$

Then, determine the number of moles of acid delivered to neutralize all the base.

$$\begin{aligned} 0.045\ M\ HCl &= \frac{0.045\ mol\ HCl}{1\ L} * 0.03075\ L = 0.00138\ mol\ HCl \\ &= 0.00138\ mol\ H_3O^+\ delivered \end{aligned}$$

This amount of acid delivered neutralizes all of the hydroxide from barium hydroxide. Thus, 0.00138 mol of OH^- existed in the original solution.



$$0.00138\ mol\ OH^- * \frac{1\ mol\ Ba(OH)_2}{2\ mol\ OH^-} = 0.00069\ mol\ Ba(OH)_2$$

$$[Ba(OH)_2] = \frac{0.00069\ mol\ Ba(OH)_2}{0.1\ L} = 0.0069\ M$$

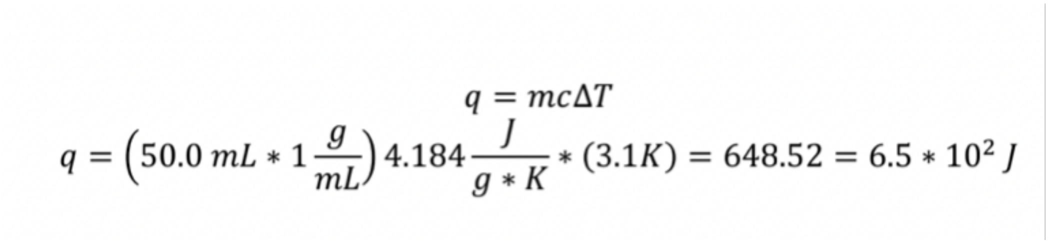
$$[OH^-] = \frac{0.00138\ mol\ OH^-}{0.1\ L} = 0.0138\ M$$

$$pOH = -\log[OH^-] = 1.86$$

$$pH = 14 - pOH = 12.14$$

6. A 12.48 g sample of an unknown metal, heated to 99°C, was then plunged into 50.0mL of 25.0°C water. The temperature of the water rose to 28.1°C. Assuming no loss of energy to the surroundings,

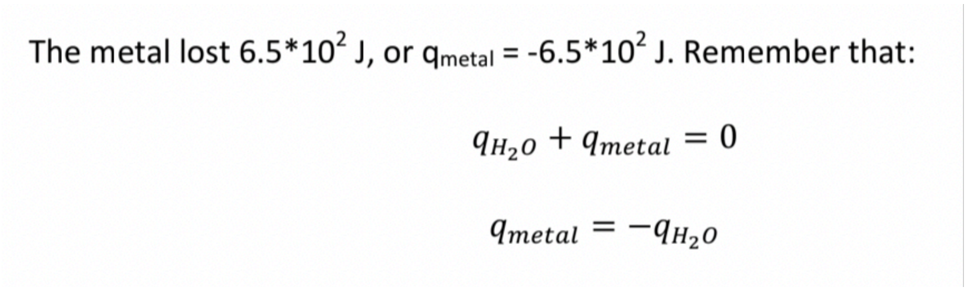
   a. How many joules of energy did the water absorb? (specific heat capacity of water = 4.184 J/(g*K))



$$q = mc\Delta T$$

$$q = \left(50.0\ mL * 1\frac{g}{mL}\right) 4.184\frac{J}{g * K} * (3.1K) = 648.52 = 6.5 * 10^2\ J$$

   b. How many joules of energy did the metal lose?

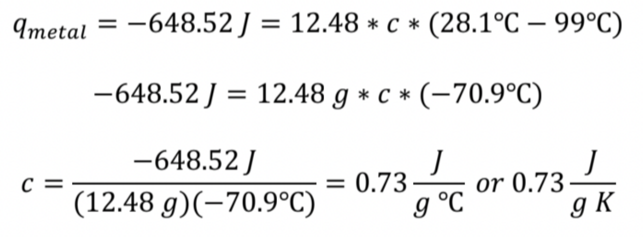


The metal lost $6.5*10^2$ J, or $q_{metal} = -6.5*10^2$ J. Remember that:

$$q_{H_2O} + q_{metal} = 0$$

$$q_{metal} = -q_{H_2O}$$

c. What is the specific heat capacity of the metal?



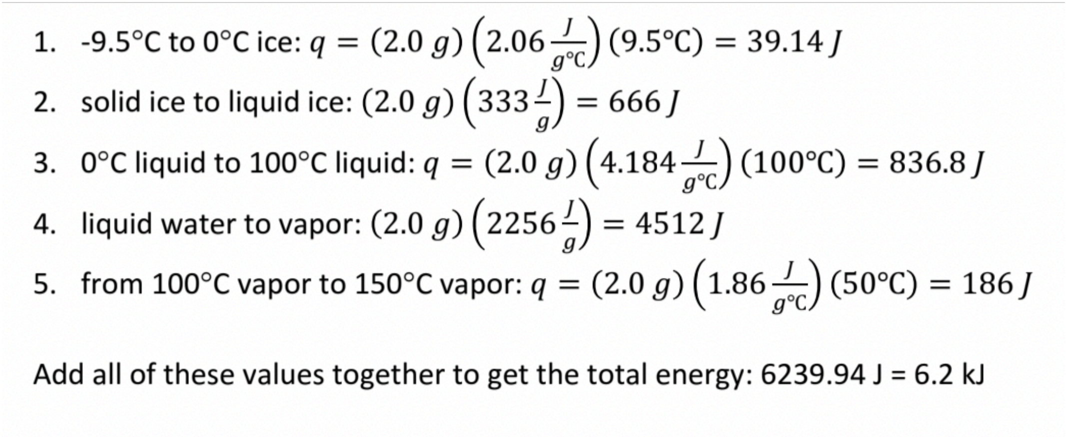
$$q_{metal} = -648.52\,J = 12.48 * c * (28.1°\text{C} - 99°\text{C})$$

$$-648.52\,J = 12.48\,g * c * (-70.9°\text{C})$$

$$c = \frac{-648.52\,J}{(12.48\,g)(-70.9°\text{C})} = 0.73\frac{J}{g\,°\text{C}}\ or\ 0.73\frac{J}{g\,K}$$

7. How much energy does it take to convert 2.0 g of ice at -9.5°C to steam at 150°C?

(Water's heat of fusion is 333 J/g; water's heat of vaporization is 2256 J/g. The specific heat capacity of: ice = 2.06 J/gK; water = 4.184 J/gK; steam = 1.86 J/gK)



1. -9.5°C to 0°C ice: $q = (2.0\,g)\left(2.06\frac{J}{g°\text{C}}\right)(9.5°\text{C}) = 39.14\,J$
2. solid ice to liquid ice: $(2.0\,g)\left(333\frac{J}{g}\right) = 666\,J$
3. 0°C liquid to 100°C liquid: $q = (2.0\,g)\left(4.184\frac{J}{g°\text{C}}\right)(100°\text{C}) = 836.8\,J$
4. liquid water to vapor: $(2.0\,g)\left(2256\frac{J}{g}\right) = 4512\,J$
5. from 100°C vapor to 150°C vapor: $q = (2.0\,g)\left(1.86\frac{J}{g°\text{C}}\right)(50°\text{C}) = 186\,J$

Add all of these values together to get the total energy: 6239.94 J = 6.2 kJ