Chapter 4: Review questions for test
p. 154-158 # 4.71, 4.99, 4.101, 4.102, 4.109, 4.112, 4.114, 4.118, 4.121, 4.132

71 How would you prepare 60.0 mL of 0.200 M HNO_3 from a stock solution of 4.00 M HNO_3?

$M_{\text{initial}}V_{\text{initial}} = M_{\text{final}}V_{\text{final}}$

You can solve the equation algebraically for V_{initial}. Then substitute in the given quantities to solve the for the volume of 4.00 *M* HNO_3 needed to prepare 60.0 mL of a 0.200 *M* HNO_3 solution.

$$\boldsymbol{V}_{\textbf{initial}} = \frac{\boldsymbol{M}_{\text{final}} \times \boldsymbol{V}_{\text{final}}}{\boldsymbol{M}_{\text{initial}}} = \frac{0.200\ M \times 60.00\ m\text{L}}{4.00\ \boldsymbol{M}} = \textbf{3.00 mL}$$

To prepare the 0.200 *M* solution, since you are starting with a relatively concentrated acid you would first add about 30 mL of water to a 60-mL volumetric flask and then you would use a graduated pipet to precisely measure 3.00 mL of the 2.00 *M* 4.00 *M* HNO_3 solution into the flask. Swirl to mix, then add water to a final volume of 60.0 mL, using a dropper to add the last amount of water just to the calibration mark. Stopper the flask and mix well. If the volume has dropped due to mixing, again use a dropper to add water to the calibration mark, stopper and mix well

99 Classify the following reactions according to the types discussed in the chapter

In redox reactions, the oxidation numbers of elements change. To test whether an equation represents a redox process, assign the oxidation numbers to each of the elements in the reactants and products. If oxidation numbers change, it is a redox reaction.

(a) $Cl_2 + 2\ OH^- \rightarrow Cl^- + ClO^- + H_2O$

On the left the oxidation number of chlorine in Cl_2 is zero (rule 1). On the right it is −1 in Cl^- (rule 2) and +1 in OCl^- (rules 3 and 5). Since chlorine is both oxidized and reduced, this is a *disproportionation redox reaction*.

(b) $Ca^{2+} + CO_3^{2-} \rightarrow CaCO_3$

The oxidation numbers of calcium and carbon do not change. This is not a redox reaction; it is a *precipitation* (double-replacement) reaction.

(c) $NH_3 + H^+ \rightarrow NH_4^+$

The oxidation numbers of nitrogen and hydrogen do not change. This is not a redox reaction; it is an *acid-base reaction*.

(d) $2\ CCl_4 + CrO_4^{2-} \rightarrow 2\ COCl_2 + CrO_2Cl_2 + 2\ Cl^-$

The oxidation numbers of carbon, chlorine, chromium, and oxygen do not change. This is not a redox reaction; it doesn't fit easily into any category, but could be considered as a type of *combination reaction*.

(e) $Ca + F_2 \rightarrow CaF_2$

The oxidation number of calcium changes from 0 to +2, and the oxidation number of fluorine changes from 0 to −1. This is a *combination redox reaction*.

(f) $2\ Li + H_2 \rightarrow 2\ LiH$: Combination Redox

(g) $Ba(NO_3)_2 + Na_2SO_4 \rightarrow 2\ NaNO_3 + BaSO_4$: Precipitation

(h) $CuO + H_2 \rightarrow Cu + H_2O$: Displacement Redox

(i) $Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$: Displacement Redox

(j) $2\ FeCl_2 + Cl_2 \rightarrow 2\ FeCl_3$: Combination Redox

101 Which of the following aqueous solutions would you expect to be the best conductor of electricity at 25°C? Explain your answer.

Choice **(d)**, 0.20 *M* $Mg(NO_3)_2$, should be the best conductor of electricity; the total ion concentration in this solution is 0.60 *M*. The total ion concentrations for solutions (a) and (c) are 0.40 *M* and 0.50 *M*, respectively. We can rule out choice (b), because acetic acid is a weak electrolyte.

102 A 5.00×10^2-mL sample of 2.00 *M* HCl solution is treated with 4.47 g of magnesium. Calculate the concentration of the acid solution after all the metal has reacted. Assume that the volume remains unchanged.

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

From the mass of Mg calculate moles of HCl reacted:

$$4.47\text{ g Mg} \times \frac{1\text{ mol Mg}}{24.31\text{ g Mg}} \times \frac{2\text{ mol HCl}}{1\text{ mol Mg}} = 0.368\text{ mol HCl reacted}$$

Next we can calculate the number of moles of HCl in the original solution.

$$5.00 \times 10^2\text{ mL} \times \frac{2.00\text{ mol HCl}}{1000\text{ mL soln}} = 1.00\text{ mol HCl}$$

Moles HCl remaining $= 1.00\text{ mol} - 0.368\text{ mol} = 0.632\text{ mol HCl}$

$$\textbf{conc. of HCl after reaction} = \frac{\text{mol HCl}}{\text{L soln}} = \frac{0.632\text{ mol HCl}}{0.500\text{ L}} = 1.26\text{ mol/L} = \mathbf{1.26}\ \boldsymbol{M}$$

109 Calculate the mass of the precipitate formed when 2.27 L of 0.0820 *M* $Ba(OH)_2$ are mixed with 3.07 L of 0.0774 *M* Na_2SO_4.

$$Ba(OH)_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + 2NaOH(aq)$$

moles $Ba(OH)_2$: $(2.27\text{ L})(0.0820\text{ mol/L}) = 0.186\text{ mol }Ba(OH)_2$
moles Na_2SO_4: $(3.06\text{ L})(0.0664\text{ mol/L}) = 0.203\text{ mol }Na_2SO_4$

Since the mole ratio between $Ba(OH)_2$ and Na_2SO_4 is 1:1, so $Ba(OH)_2$ is the limiting reagent. The mass of $BaSO_4$ formed is:

$$0.186\text{ mol Ba(OH)}_2 \times \frac{1\text{ mol BaSO}_4}{1\text{ mol Ba(OH)}_2} \times \frac{233.4\text{ g BaSO}_4}{\text{mol BaSO}_4} = \mathbf{43.4\ g\ BaSO_4}$$

112 A 1.00-g sample of a metal X (that is known to form X^{2+} ions) was added to a 0.100 L of 0.500 *M* H_2SO_4. After all the metal had reacted, the remaining acid required 0.0334 L of 0.500 *M* NaOH solution for neutralization. Calculate the molar mass of the metal and identify the element.

The balanced equations for the two reactions are:

$$X(s) + H_2SO_4(aq) \longrightarrow XSO_4(aq) + H_2(g)$$

$$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

First, let's find the number of moles of excess acid from the reaction with NaOH.

$$0.0334\text{ L} \times \frac{0.500\text{ mol NaOH}}{1\text{ L soln}} \times \frac{1\text{ mol H}_2\text{SO}_4}{2\text{ mol NaOH}} = 8.35 \times 10^{-3}\text{ mol H}_2\text{SO}_4$$

The original number of moles of acid was:

$$0.100\text{ L} \times \frac{0.500\text{ mol } H_2SO_4}{1\text{ L soln}} = 0.0500\text{ mol } H_2SO_4$$

The amount of sulfuric acid that reacted with the metal, X, is

$$(0.0500\text{ mol } H_2SO_4) - (8.35 \times 10^{-3}\text{ mol } H_2SO_4) = 0.0417\text{ mol } H_2SO_4.$$

Since the mole ratio from the balanced equation is 1 mole X : 1 mole H_2SO_4, then the amount of X that reacted is 0.0417 mol X.

$$\textbf{molar mass X} = \frac{1.00\text{ g X}}{0.0417\text{ mol X}} = \mathbf{24.0\ g / mol}$$

The element is **magnesium**.

114 **A 60.0-mL 0.513 *M* glucose ($C_6H_{12}O_6$) solution is mixed with 120.0 mL of 2.33 *M* glucose solution. What is the concentration of the final solution? Assume the volumes are additive.**
First, calculate the number of moles of glucose present.

$$\frac{0.513\text{ mol glucose}}{1000\text{ mL soln}} \times 60.0\text{ mL} = 0.0308\text{ mol glucose}$$

$$\frac{2.33\text{ mol glucose}}{1000\text{ mL soln}} \times 120.0\text{ mL} = 0.280\text{ mol glucose}$$

Add the moles of glucose, then divide by the total volume of the combined solutions to calculate the molarity.

$$60.0\text{ mL} + 120.0\text{ mL} = 180.0\text{ mL} = 0.180\text{ L}$$

$$\textbf{Molarity of final solution} = \frac{(0.0308 + 0.280)\, mol\text{ glucose}}{0.180\ L} = 1.73\text{ mol/L} = \mathbf{1.73}\ \boldsymbol{M}$$

118 **Using the apparatus shown in Figure 4.1, a student found that a sulfuric acid solution caused the light bulb to glow brightly. However, after the addition of a certain amount of barium hydroxide [$Ba(OH)_2$] solution, the light began to dim even though $Ba(OH)_2$ is also a strong electrolyte. Explain.**
Since both of the original solutions were strong electrolytes, you would expect a mixture of the two solutions to also be a strong electrolyte. However, since the light dims, the mixture must contain fewer ions than the original solution. Indeed, H^+ from the sulfuric acid reacts with the OH^- from the barium hydroxide to form water. The barium cations react with the sulfate anions to form insoluble barium sulfate.

$$2H^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2OH^-(aq) \longrightarrow 2H_2O(l) + BaSO_4(s)$$

Thus, the reaction depletes the solution of ions and the conductivity decreases.

121 The concentration of lead ions (Pb^{2+}) in a sample of polluted water that also contains NO_3^- ions is determined by adding solid Na_2SO_4 to exactly 500. mL of the water.

(a) Write the molecular and net ionic equation for the reaction.

$$Pb(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow PbSO_4(s) + 2NaNO_3(aq)$$

$$Pb^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s)$$

(b) Calculate the molar concentration of Pb^{2+} if 0.00450 g of Na_2SO_4 was needed for the complete precipitation of Pb^{2+} as $PbSO_4$.

$$0.00450 \text{ g } Na_2SO_4 \times \frac{1 \text{ mol } Na_2SO_4}{142.1 \text{ g } Na_2SO_4} \times \frac{1 \text{ mol } Pb(NO_3)_2}{1 \text{ mol } Na_2SO_4} \times \frac{1 \text{ mol } Pb^{2+}}{1 \text{ mol } Pb(NO_3)_2} = 3.17 \times 10^{-5} \text{ mol } Pb^{2+}$$

$$\mathbf{[Pb^{2+}]} = \frac{\text{mol } Pb^{2+}}{\text{L of soln}} = \frac{3.17 \times 10^{-5} \text{ mol } Pb^{2+}}{0.500 \text{ L soln}} = \boxed{\mathbf{6.34 \times 10^{-5}}\ \boldsymbol{M}}$$

132 A useful application of oxalic acid is the removal of rust (Fe_2O_3) from bathtub rings according to the reaction

$$\mathbf{Fe_2O_3(s) + 6H_2C_2O_4(aq) \rightarrow 2Fe(C_2O_4)_3^{3-} + 3H_2O + 6H^+(aq)}$$

Calculate the number of grams of rust that can be removed by 5.00×10^2 mL of a 0.100 *M* solution of oxalic acid.

$$? \text{ mol } H_2C_2O_4 = 5.00 \times 10^2 \text{ mL} \times \frac{0.100 \text{ mol } H_2C_2O_4}{1000 \text{ mL soln}} = 0.0500 \text{ mol } H_2C_2O_4$$

$$? \text{ g } Fe_2O_3 = 0.0500 \text{ mol } H_2C_2O_4 \times \frac{1 \text{ mol } Fe_2O_3}{6 \text{ mol } H_2C_2O_4} \times \frac{159.7 \text{ g } Fe_2O_3}{1 \text{ mol } Fe_2O_3} = \boxed{\mathbf{1.33 \text{ g } Fe_2O_3}}$$

Predicting Products For each of the following situations, assume that a reaction takes place and write out the complete molecular equation. Also write a net ionic equation if one can be written. Then, determine if a reaction would actually take place and explain how you know.

A. Aqueous lead(II) nitrate is added to aqueous aluminum chloride

Mol. Eq: $3\ Pb(NO_3)_2\ (aq) + 2\ AlCl_3\ (aq) \rightarrow 3\ PbCl_2\ (s) + 2\ Al(NO_3)_3\ (aq)$

Net Ionic Eq: $Pb^{2+} + 2\ Cl^- \rightarrow PbCl_2$

Does this reaction occur? ____*Yes*____ *Explain how you know.*
$PbCl_2$ is insoluble, so the precipitation reaction occurs.

B. Calcium metal is added to water.

Mol. Eq: $Ca\ (s) + 2\ H_2O\ (l) \rightarrow Ca(OH)_2\ (aq) + H_2\ (g)$

Net Ionic Eq: $Ca + H_2O \rightarrow Ca^{2+} + 2\ OH^- + H_2$

Does this reaction occur? ____*Yes*____ *Explain how you know. Any of the following is good:*

- Ca is easier to oxidize than H_2 since it has a more negative reduction potential
- H_2O is easier to reduce than Ca^{2+} since it has a more positive reduction potential

C. Silver wire is immersed in aqueous sulfuric acid.

Mol. Eq: $2\ Ag\ (s) + H_2SO_4\ (aq) \rightarrow Ag_2SO_4\ (aq) + H_2\ (g)$

Net Ionic Eq: $2\ Ag + 2\ H^+ \rightarrow 2\ Ag^+ + H_2$

Does this reaction occur? ____*No*____ *Explain how you know.*

- Ag is harder to oxidize than H_2 since it has a more positive RP
- H^+ is harder to reduce than Ag^+ since it has a more negative RP

D. Chlorine gas is bubbled through an aqueous solution of potassium fluoride.

Mol. Eq: $Cl_2\ (g) + 2\ KF\ (aq) \rightarrow 2\ KCl\ (aq) + F_2\ (g)$

Net Ionic Eq: $Cl_2 + 2\ F^- \rightarrow 2\ Cl^- + F_2$

Does this reaction occur? ____*No*____ *Explain how you know.*
Cl_2 is harder to reduce than F_2 since it has a more negative RP
F^- is harder to oxidize than Cl^- since it has a more positive RP