Chem 2 AP Homework #4-5: Acid Base and Redox Titrations KEY
(Take from book, Problems pg. 155 #82, 83, 84, 86, 88, 90, 92, 94, 96)

4.82 **Why does an acid-base indicator change color when the pH changes? (What happens to molecule?)**
The indicator changes color when a proton is either removed or added.

4.83 **A student carried out two titrations to standardize a NaOH solution of unknown concentration by titrating the NaOH with KHP (potassium hydrogen phlalate-an acid). She titrated the NaOH with two similar masses of KHP but one time the KHP was dissolved in <u>20.00 mL of distilled water</u> and the other time the KHP was dissolved in <u>40.00 mL of water</u>. Assuming no experimental error, would she obtain the same result for the concentration of the NaOH solution both times?**
<u>Yes</u>, she should obtain the same result for the concentration of the NaOH solution. The amount of NaOH needed should be proportional to the moles of acid, which is known for both titrations, and independent of the amount of solvent present.

4.84 **Would the volume of a 0.20 *M* NaOH solution needed to titrate 25.0 mL of a 0.10 *M* HNO_2 (a weak acid) solution to its equivalence point be different from the volume of NaOH needed to titrate 25.0 mL of a 0.10 *M* HCl (a strong acid) solution? The data is organized below:**

		or		
NaOH +	**HNO_2 →**	**or**	**NaOH +**	**HCl →**
0.20M	**0.10M**		**0.20M**	**0.10M**
	25 mL			**25 mL**

<u>No</u>, the volumes would not be different. The equivalence point is the point at which the moles of OH^- is equal to the moles of H^+. In this case, the moles of H^+ are the same for both of the acids. And since NaOH is a strong base, it drives both reactions 100% to the right, so all moles of both acids react.

4.86 **Calculate the concentration (in molarity) of a NaOH solution if 25.0 mL of the solution are needed to neutralize 17.4 mL of a 0.312 *M* HCl solution.**

First, complete the molecular eq: HCl(*aq*) + NaOH(*aq*) → $H_2O(l)$ + NaCl(*aq*)

$$?\ mol\ NaOH = 17.4\ mL\ HCl \times \frac{0.312\ mol\ HCl}{1000\ mL\ \text{soln}} \times \frac{1\ mol\ NaOH}{1\ mol\ HCl} = 5.43 \times 10^{-3}\ mol\ NaOH$$

$$\textbf{\textit{M} of NaOH} = \frac{5.43 \times 10^{-3}\ mol\ NaOH}{25.0\ mL\ \text{soln}} \times \frac{1000\ \text{mL}}{1\ \text{L}} = \boxed{\mathbf{0.217}\ \textbf{\textit{M}}}$$

4.88 **What volume of a 0.500 *M* HCl solution is needed to neutralize each of the following?**
(a) 10.0 mL of a 0.300 *M* NaOH solution

Molecular: HCl(*aq*) + NaOH(*aq*) ⟶ NaCl(*aq*) + $H_2O(l)$

$$10.0\ mL \times \frac{0.300\ mol\ NaOH}{1000\ mL\ of\ solution} \times \frac{1\ mol\ HCl}{1\ mol\ NaOH} = 3.00 \times 10^{-3}\ mol\ HCl$$

$$\textbf{volume of HCl} = 3.00 \times 10^{-3}\ mol\ HCl \times \frac{1000\ \text{mL}}{0.500\ \text{mol}} = \boxed{\mathbf{6.00\ mL}}$$

(b) 10.0 mL of a 0.200 *M* $Ba(OH)_2$ solution

Molelcular: 2 HCl(*aq*) + $Ba(OH)_2(aq)$ ⟶ $BaCl_2(aq)$ +2$H_2O(l)$

$$10.0\ mL \times \frac{0.200\ mol\ Ba(OH)_2}{1000\ mL\ of\ solution} \times \frac{2\ mol\ HCl}{1\ mol\ Ba(OH)_2} = 4.00 \times 10^{-3}\ mol\ HCl$$

$$\textbf{volume of HCl} = 4.00 \times 10^{-3}\ mol\ HCl \times \frac{1000\ \text{mL}}{0.500\ \text{mol}} = \boxed{\mathbf{8.00\ mL}}$$

4.90 Explain why potassium permanganate ($KMnO_4$) and potassium dichromate ($K_2Cr_2O_7$) can serve as internal indicators in redox titrations.

MnO_4^- and $Cr_2O_7^{2-}$ both change colors when reduced, so their color change acts as an indicator.

MnO_4^- (purple) → Mn^{+2} (colorless) $Cr_2O_7^{2-}$ (orange)→ Cr^{3+} (green)

4.92 The SO_2 present in air is mainly responsible for the acid rain phenomenon. Its concentration can be determined by titrating against a standard permanganate solution as follows.

$$5SO_2 + 2MnO_4^- + 2H_2O \longrightarrow 5SO_4^{2-} + 2Mn^{2+} + 4H^+$$

Calculate the number of grams of SO_2 in a sample of air if 7.37 mL of 0.00800 M $KMnO_4$ solution are required for the titration.

$$\frac{0.00800\ mol\ KMnO_4}{1000\ mL\ \text{soln}} \times 7.37\ mL = 5.90 \times 10^{-5}\ mol\ KMnO_4$$

$$(5.90 \times 10^{-5}\ mol\ KMnO_4) \times \frac{5\ mol\ SO_2}{2\ mol\ KMnO_4} \times \frac{64.07\ g\ SO_2}{1\ mol\ SO_2} = \mathbf{9.45 \times 10^{-3}\ g\ SO_2}$$

4.94 The concentration of a hydrogen peroxide solution can be conveniently determined by titration against a standardized potassium permanganate solution in an acidic medium according to the following equation:

$$2MnO_4^- + 5H_2O_2 + 6H^+ \longrightarrow 5O_2 + 2Mn^{2+} + 8H_2O$$

If 36.44 mL of a 0.01652 M $KMnO_4$ solution are required to oxidize 25.00 mL of a H_2O_2 solution, calculate the molarity of the H_2O_2 solution.

$$\frac{0.01652\ mol\ KMnO_4}{1000\ mL\ \text{soln}} \times 36.44\ mL = 6.020 \times 10^{-4}\ mol\ KMnO_4$$

$$(6.020 \times 10^{-4}\ mol\ KMnO_4) \times \frac{5\ mol\ H_2O_2}{2\ mol\ KMnO_4} = 1.505 \times 10^{-3}\ mol\ H_2O_2$$

$$\textbf{Molarity of } \mathbf{H_2O_2} = \frac{1.505 \times 10^{-3}\ mol\ H_2O_2}{0.02500\ L} = \mathbf{0.06020}\ \boldsymbol{M}$$

4.96 Oxalic acid ($H_2C_2O_4$) is present in many plants and vegetables. If 24.0 mL of 0.0100 M $KMnO_4$ solution is needed to titrate 1.00 g of a sample of $H_2C_2O_4$(oxalic acid) to the equivalence point, what is the percent by mass of $H_2C_2O_4$ in the sample?

Net Ionic Eq: $2MnO_4^- + 6H^+ + 5\ H_2C_2O_4 \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

$$\frac{0.0100\ mol\ KMnO_4}{1000\ mL\ \text{soln}} \times 24.0\ mL = 2.40 \times 10^{-4}\ mol\ KMnO_4$$

$$(2.40 \times 10^{-4}\ mol\ KMnO_4) \times \frac{5\ mol\ H_2C_2O_4}{2\ mol\ KMnO_4} \times \frac{90.04\ g\ H_2C_2O_4}{1\ mol\ H_2C_2O_4} = 0.0540\ g\ H_2C_2O_4$$

$$\textbf{mass \%} = \frac{0.0540\ g}{1.00\ g} \times 100\% = \mathbf{5.40\%\ H_2C_2O_4}$$