ANSWER KEY

Acid-Base Neutralization & Titrations Practice

Name:______________________

1. A 35.0 mL sample of NaOH solution is titrated to an end point by 14.8 mL 0.412*M* HBr solution. What is the molarity of the NaOH solution?

   __1__ HBr + __1__ NaOH → __1__ NaBr + __1__ H_2O

$$M_AV_A = M_BV_B \rightarrow M_B = \frac{M_AV_A}{V_B} \rightarrow M_B = \frac{(0.412M)(14.8mL)}{(35.0mL)} \rightarrow \boxed{M_B = 0.174\ M\ NaOH}$$

2. If 25.0 mL of KOH were needed to neutralize 15.0 mL of 3.50*M* HBr, calculate the molarity of the base.

   __1__ HBr + __1__ KOH → __1__ KBr + __1__ H_2O

$$M_AV_A = M_BV_B \rightarrow M_B = \frac{M_AV_A}{V_B} \rightarrow M_B = \frac{(3.50M)(15.0mL)}{(25.0mL)} \rightarrow \boxed{M_B = 2.10\ M\ KOH}$$

3. It required 34.0 mL of 1.90*M* KOH to neutralize 25.0 mL of H_3PO_4. Calculate the molarity of the phosphoric acid.

   __1__ H_3PO_4 + __3__ KOH → __1__ K_3PO_4 + __3__ H_2O

$$\frac{M_AV_A}{n_A} = \frac{M_BV_B}{n_B} \rightarrow M_A = \frac{M_BV_Bn_A}{V_An_B} \rightarrow M_A = \frac{(1.90M)(34.0mL)(1)}{(25.0mL)(3)} \rightarrow \boxed{M_A = 0.860\ M\ H_3PO_4}$$

4. A student titrated 20.0 mL of 0.500*M* HNO_3 with NaOH. Her data is shown in the following table.

   a. Complete the following table:

	Trial 1	Trial 2	Trial 3
INITIAL VOLUME NaOH	0.00 mL	15.3 mL	30.4 mL
FINAL VOLUME NaOH	15.3 mL	30.4 mL	45.3 mL
VOLUME NaOH used	15.3mL	15.1mL	14.9mL

   b. Calculate the average volume of NaOH used from the three trials.

$$\frac{(15.3mL)+(15.1mL)+(14.9mL)}{3} \rightarrow \frac{45.3\ mL}{3} \rightarrow \boxed{\text{Avg Volume} = 15.1\ mL\ NaOH}$$

   c. Use your average volume of NaOH used to calculate the molarity of the base.

   __1__ HNO_3 + __1__ NaOH → __1__ $NaNO_3$ + __1__ H_2O

$$M_AV_A = M_BV_B \rightarrow M_B = \frac{M_AV_A}{V_B} \rightarrow M_B = \frac{(0.500M)(20.0mL)}{(15.1mL)} \rightarrow \boxed{M_B = 0.662M\ NaOH}$$

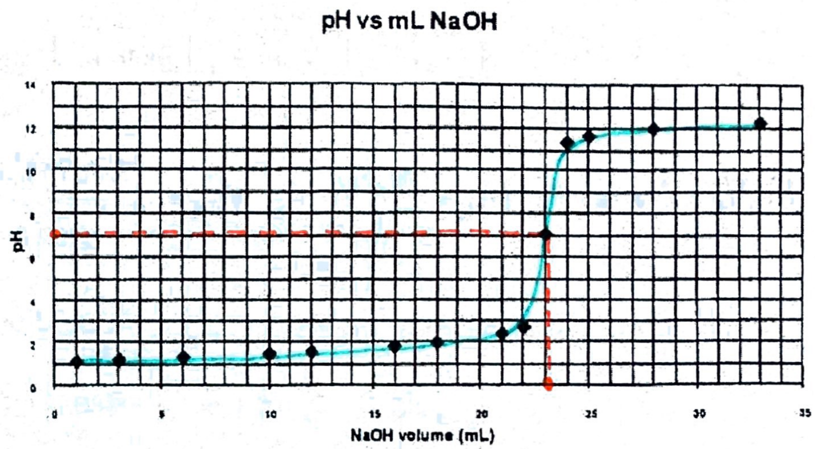
5. What is the molarity of a nitric acid solution if 43.3 mL of 0.100*M* potassium hydroxide solution is needed to neutralize 20.0 mL of the unknown concentration of nitric acid solution?

$\underline{1}\ HNO_3 + \underline{1}\ KOH \rightarrow \underline{1}\ KNO_3 + \underline{1}\ H_2O$

$M_AV_A = M_BV_B \rightarrow M_A = \frac{M_BV_B}{V_A} \rightarrow M_A = \frac{(0.100\ M)(43.3\ mL)}{(20.0\ mL)} \rightarrow \boxed{M_A = 0.217\ M\ HNO_3}$

Titration Problems

1) You have a solution of HBr of unknown concentration. To figure out the concentration, you measure out 25 mL of the acid solution. You then begin adding 0.100M NaOH solution and measuring the pH after each addition. The graph of pH vs mL NaOH looks like this:



a) At what volume of NaOH solution was the HBr consumed? (This is called the equivalence point)
Equivalence Point → $\boxed{pH = 7}$ & $\boxed{23\ mL}$

b) Write a balanced equation for the reaction.
$\underline{1}\ HBr + \underline{1}\ NaOH \rightarrow \underline{1}\ NaBr + \underline{1}\ H_2O$ → SA + SB → Neutral ∴ pH = 7

c) Calculate the moles of acid that must have been in the original 25 mL of acid.
1:1 Ratio ∴ @ Equivalence Point mol $[H^+]$ = mol $[OH^-]$ →

$M_B = \frac{\text{mol solute}_B}{\text{L soln}_B} \rightarrow \text{mol solute}_B = (M_B)(L_B)$

mol NaOH = (0.100 M NaOH)(0.023 L NaOH) =

mol NaOH = $\boxed{0.0023\ \text{mol NaOH}}$

$\boxed{0.0023\ \text{mol HBr}}$

d) Calculate the concentration of the original acid solution.

$M_A = \frac{\text{mol solute A}}{\text{L soln A}} \rightarrow M_A = \frac{0.0023\ \text{mol HBr}}{0.025\ \text{L soln}}$

$\boxed{M_A = 0.0920\ M\ HBr}$