AP CHEMISTRY

TOPIC 7: ACIDS & BASES, PART F, PART II

Day 94:

- Titrations
- pH curve
- Indicators

$$K_a = \frac{\left[H_3O^{+1} \right]\left[OBr^{-1} \right]}{\left[HOBr \right]} = 2.00\times10^{-9}$$

Hypobromous acid, HOBr, is a weak acid in water. The K_a expression for HOBr is shown above.

a) Write a chemical equation showing how HOBr behaves as an acid in water.

$$HOBr + H_2O \rightleftharpoons H_3O^{+1} + OBr^{-1} \quad \textbf{\textit{OR}} \quad HOBr \rightleftharpoons H^{+1} + OBr^{-1}$$

b) Calculate the pH of a 0.175 *M* solution of HOBr.

	[HOBr]	$\rightleftharpoons$	[OBr^{-1}]	+	[H^{+1}]
I	0.175		0		0
C	- x		+ x		+ x
E	0.175 - x		x		x

$$K_a = \frac{\left[H^{+1} \right]\left[OBr^{-1} \right]}{\left[HOBr \right]} = \frac{x^2}{\left(0.175 - x \right)} = \frac{x^2}{\left(0.175 \right)} = 2.00\times10^{-9}$$

$$x^2 = \left(2.00\times10^{-9} \right)\left(0.175 \right) = 3.50\times10^{-10}$$

$$x = \sqrt{3.50\times10^{-10}} = 1.87\times10^{-5}$$

$$x = [\, H_3O^{+1}\,] = [\, H^{+1}\,] = 1.87 \text{ x } 10^{-5}\ M$$

$$\textbf{pH} = -\log\left(1.87 \text{ x } 10^{-5} \right) = \mathbf{4.73}$$

c) Write the net ionic equation for the reaction between the weak acid, $HOBr_{(aq)}$ and the strong base $NaOH_{(aq)}$.

$$NaOH \rightarrow Na^{+1} + OH^{-1} \qquad\qquad HOBr \rightleftharpoons H^{+1} + OBr^{-1}$$

Ionic Equation:

$$Na^{+1}_{(aq)} + OH^{-1}_{(aq)} + H^{+1}_{(aq)} + OBr^{-1}_{(aq)} \rightarrow Na^{+1}_{(aq)} + OBr^{-1}_{(aq)} + \boldsymbol{H_2O_{(l)}}$$

Net Ionic Equation:

$$OH^{-1}_{(aq)} + H^{+1}_{(aq)} \rightarrow H_2O_{(l)}$$

Net Ionic Equation (Definition): A net ionic equation is written by dropping out the spectator ions and showing only those chemical species that are involved in the chemical reaction.

This reaction is an important (**neutralization**) reaction that occurs when an acid (H^{+1}) reacts with a base (OH^{-1}).

d) In an experiment, 20.00 mL of 0.175 *M* HOBr $_{(aq)}$ is placed in a flask and titrated with 6.55 mL of 0.435 *M* NaOH $_{(aq)}$.

(i) Calculate the number of moles of NaOH $_{(aq)}$ added.

$$\text{Number of moles (NaOH):} \;=\; \frac{6.55\ mL}{} \times \frac{1\ L}{1000\ mL} \times \frac{0.435\ mol}{L} \;=\; 0.00285\ mol\ NaOH \;=\; OH^{-1}$$

(ii) Calculate [H_3O^{+1}] (or [H^{+1}]) in the flask after the NaOH $_{(aq)}$ has been added.

$$\text{Number of moles (HOBr):} \;=\; \frac{20.00\ mL}{} \times \frac{1\ L}{1000\ mL} \times \frac{0.175\ mol}{L} \;=\; 0.00350\ mol\ HOBr$$

Total Volume: 20.0 mL + 6.55 mL = 26.55 mL = 0.02655 L

$$\left[\ HOBr\ \right] = \frac{0.00350\ mol}{0.02655\ L} = 0.1318\ M \quad ; \quad \left[\ OH^{-1}\ \right] = \frac{0.00285\ mol}{0.02655\ L} = 0.107\ M$$

	OBr^{-1}	+	H_2O	$\rightleftharpoons$	HOBr	+	OH^{-1}
I	~ 0		-		0.1318		0.107
C	+ 0.107		-		- 0.107		- 0.107
E	0.107		-		0.0248		~ 0

$$K_b = \frac{K_w}{K_a} = \frac{1.00\times10^{-14}}{2.00\times10^{-9}} = 5.00\times10^{-6}$$

$$K_b = \frac{\left[\ HOBr\ \right]\left[\ OH^{-1}\ \right]}{\left[\ OBr^{-1}\ \right]} = \frac{\left(\ 0.0248\ \right)\left[\ OH^{-1}\ \right]}{\left(\ 0.107\ \right)} = 5.00\times10^{-6}$$

$$\frac{\left(\ 5.00\times10^{-6}\ \right)\left(\ 0.107\ \right)}{\left(\ 0.0248\ \right)} = 2.15\times10^{-5} = \left[\ OH^{-1}\ \right]$$

$$K_w = \left[\ H^{+1}\ \right]\left[\ OH^{-1}\ \right] = 1.00\times10^{-14}$$

$$\left[\ H^{+1}\ \right] = \frac{1.00\times10^{-14}}{\left[\ OH^{-1}\ \right]} = \frac{1.00\times10^{-14}}{2.15\times10^{-5}} = 4.64\times10^{-10}$$

OR, you could find the pH and then calculate the Antilog (- pH) to get the [H^{+1}].

(iii) Calculate [OH^{-1}] in the flask after the NaOH $_{(aq)}$ has been added.

$$K_w = \left[\ H^{+1}\ \right]\left[\ OH^{-1}\ \right] = 1.00\times10^{-14}$$

$$\left[\ OH^{-1}\ \right] = \frac{1.00\times10^{-14}}{\left[\ H^{+1}\ \right]} = \frac{1.00\times10^{-14}}{4.64\times10^{-10}} = 2.15\times10^{-5}\ M$$

(iv) Calculate the pH of the titration solution after the NaOH $_{(aq)}$ has been added.

$$\textbf{pH} = \textbf{-}\log\ (\ 4.64 \text{ x } 10^{-10}\ ) = \textbf{9.33}$$

(v) Is the titration solution at the equivalence point? Justify your answer.

No, At the equivalence point, the number of moles of HOBr (acid) will be equal to the number of moles of OH^{-1} (base).