AP CHEMISTRY

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

Day 94:

- Titrations
- pH curve
- Indicators

$$K_b = \frac{\left[\ OH^{-1}\ \right]\left[\ CH_4OH_2^{\ +1}\ \right]}{\left[\ CH_4OH\ \right]} = 4.57 \times 10^{-5}$$

Methylamine, CH_4OH, is a weak base in water. The K_b expression is shown above.

a) Write a chemical equation showing how CH_4OH behaves as an base in water.

$$CH_4OH + HOH \rightleftharpoons OH + CH_4OH_2^{\ +1}$$

b) Calculate the pH of a 0.588 *M* solution of CH_4OH.

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

$$K_a = \frac{\left[\ H^{+1}\ \right]\left[\ OBr^{-1}\ \right]}{\left[\ HOBr\ \right]} = \frac{x^2}{(\ 0.175 - x\ )} = \frac{x^2}{(\ 0.175\ )} = 2.00 \times 10^{-9}$$

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

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

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

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

c) Write the net ionic equation for the reaction between the weak base, $CH_4OH_{(aq)}$ and the strong acid, $HNO_{3\ (aq)}$.

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

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, 42.0 mL of 0.588 M $CH_4OH_{(aq)}$ is placed in a flask and titrated with 10.0 mL of 0.400 M $HNO_{3\,(aq)}$.

(i) Calculate the number of moles of $HNO_{3\,(aq)}$ 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 [OH^{-1}] in the flask after the $HNO_{3\,(aq)}$ 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 [H^{+1}] in the flask after the $HNO_{3\ (aq)}$ has been added.

$$K_w = \left[H^{+1} \right] \left[OH^{-1} \right] = 1.00 \times 10^{-14}$$

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

(iv) Calculate the pH of the titration solution after the $HNO_{3\ (aq)}$ has been added.

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