Chemistry Lecture #95: Ion Product Constant for Water & pH

Water molecules can ionize to form $H^+$ and $OH^-$ in an equilibrium reaction. From the equilibrium reaction, we write an equilibrium constant expression.

$$H_2O \rightleftharpoons H^+ + OH^-$$

$$Keq = \frac{[H^+][OH^-]}{[H_2O]}$$

Very few water molecules ionize to form $H^+$ and $OH^-$. Thus, the concentration of water, $[H_2O]$, changes very little and can be considered a constant number. Multiplying both sides by $[H_2O]$,

$$Keq \times [H_2O] = \frac{[H^+][OH^-] \times [H_2O]}{[H_2O]} = 1$$

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$

$K_w$ is the ion product constant for water. It is equal to $1.0 \times 10^{-14}$ when the temperature is 25 degrees Celsius.

$K_w$ can be used to find the concentration of $H^+$ or $OH^-$. 
The concentration of hydronium ion in an aqueous solution is $2.0 \times 10^{-3}$ M. What is the concentration of hydroxide?

\[
\text{Kw} = [\text{H}^+][\text{OH}^-] \\
1.0 \times 10^{-14} = [2.0 \times 10^{-3}][\text{OH}^-]
\]

\[
\frac{1.0 \times 10^{-14}}{2.0 \times 10^{-3}} \frac{[2.0 \times 10^{-3}]}{[2.0 \times 10^{-3}]} \\
[\text{OH}^-] = 5.0 \times 10^{-2} \text{ M}
\]

All aqueous (water based) solutions contain H$^+$ and OH$^-$. If the amount of H$^+$ is greater than the amount of OH$, the solution is acidic. If the amount of OH$^-$ is greater than the amount of H$^+$, the solution is basic. If the amounts of H$^+$ and OH$^-$ are equal, the solution is neutral.

One way to describe the relative amount of H$^+$ in solution is to use pH. Mathematically,

\[
\text{pH} = -\log[\text{H}^+]
\]

Quick review: \(\log(100) = 2\), \(\log(1000) = 3\), \(\log(0.0001) = -4\). \(10^2 = 100\), \(10^3 = 1000\), \(10^{-4} = 0.0001\)
The pH of an aqueous solution with a hydronium ion concentration of $7.4 \times 10^{-3}$ would be

\[
\text{pH} = -\log[H^+]
\]
\[
\text{pH} = -\log[7.4 \times 10^{-3}]
\]
\[
\text{pH} = -(2.13)
\]
\[
\text{pH} = 2.13
\]

Notice that the pH is written with two numbers past the decimal: .13. For pH, the number of places past the decimal that you write equals the number of significant figures in the $H^+$ concentration. $7.4 \times 10^{-3}$ has two significant figures, so we write the pH with two numbers to the right of the decimal point.

If the pH of a solution is less than 7, the solution is acidic. Thus, a pH = 2.13 tells us that we have an acidic solution.

If the pH of a solution is greater than 7, the solution is basic.

If the pH of a solution is equal to 7, the solution is neutral.

Knowing the pH of a solution, we can calculate $[H^+]$.

Find $[H^+]$ if the pH is 8.20

\[
\text{pH} = -\log[H^+]
\]
\[
8.20 = -\log[H^+]
\]
\[
-8.20 = \log \left[ H^+ \right]
\]
\[
\left[ H^+ \right] = 10^{-8.20}
\]
\[
\left[ H^+ \right] = 6.3 \times 10^{-9} \text{ M}
\]
If the pH of an acid solution is known, we can calculate $K_a$.

Find the $K_a$ for a 0.0400 M solution of HClO$_2$ (chlorous acid) if its pH is 1.80.

Answer

First, find $[H^+]$.

$$\text{pH} = -\log[H^+]$$
$$1.80 = -\log[H^+]$$
$$[H^+] = 10^{-1.8}$$
$$[H^+] = 0.016 \text{ M}$$

This value also represents the number of HClO$_2$ molecules that ionized from the original 0.0400 M that we started with.

$$\begin{align*}
0.0400 - 0.016 & \quad 0.016 & \quad 0.016 \\
\text{HClO}_2 & \quad \rightleftharpoons & \quad \text{H}^+ & \quad + & \quad \text{ClO}_2^- \\
\end{align*}$$

$$K_a = \frac{[H^+][\text{ClO}_2^-]}{[\text{HClO}_2]}$$

$$K_a = \frac{0.016\times0.016}{0.0400 - 0.016}$$

$$K_a = 1.1 \times 10^{-2}$$