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.

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

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

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 $[H_2O]$,

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

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

Kw is the ion product constant for water. It is equal to 1.0×10^{-14} when the temperature is 25 degrees Celsius.

Kw can be used to find the concentration of H^+ or OH^-.

The concentration of hydronium ion in an aqueous solution is 2.0×10^{-3} M. What is the concentration of hydroxide?

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

$$\frac{1.0 \times 10^{-14}}{2.0 \times 10^{-3}} = \frac{[2.0 \times 10^{-3}][OH^-]}{[2.0 \times 10^{-3}]}$$

$[OH^-] = 5.0 \times 10^{-12}$ 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,

$$pH = -\log[H^+]$$

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 7.4×10^{-3} would be

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

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. 7.4×10^{-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

$pH = -\log[H^+]$
$8.20 = -\log[H^+]$
$-8.20 = \log[H^+]$
$[H^+] = 10^{-8.20}$
$[H^+] = 6.3 \times 10^{-9}$ M

If the pH of an acid solution is known, we can calculate Ka.

Find the Ka for a 0.0400 M solution of $HClO_2$ (chlorous acid) if its pH is 1.80.

Answer

First, find $[H^+]$.

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

This value also represents the number of $HClO_2$ molecules that ionized from the original 0.0400 M that we started with.

$$\begin{array}{cccc} 0.0400 - 0.016 & & 0.016 & 0.016 \\ HClO_2 & \rightleftharpoons & H^+ & + \; ClO_2^- \end{array}$$

$$Ka = \frac{[H^+][ClO_2^-]}{[HClO_2]}$$

$$Ka = \frac{[0.016][0.016]}{[0.0400 - 0.016]}$$

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