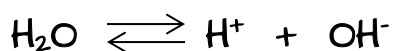


## Chemistry Lecture #95: Ion Product Constant for Water & pH

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



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

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  $[H_2O]$ ,

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

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

$K_w$  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?

$$K_w = [H^+][OH^-]$$

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

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

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

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,

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

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  $7.4 \times 10^{-3}$  would be

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

$$\text{pH} = -\log[7.4 \times 10^{-3}]$$

$$\text{pH} = -(-2.13)$$

$$\text{pH} = 2.13$$

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  $\text{H}^+$  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  $[\text{H}^+]$ .

Find  $[\text{H}^+]$  if the pH is 8.20

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

$$8.20 = -\log[\text{H}^+]$$

$$-8.20 = \log [\text{H}^+]$$

$$[\text{H}^+] = 10^{-8.20}$$

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

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  $\text{HClO}_2$  (chlorous acid) if its pH is 1.80.

*Answer*

First, find  $[\text{H}^+]$ .

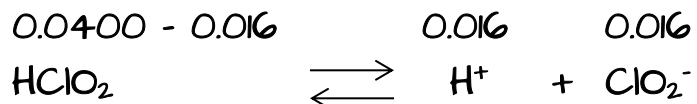
$$\text{pH} = -\log[\text{H}^+]$$

$$1.80 = -\log[\text{H}^+]$$

$$[\text{H}^+] = 10^{-1.8}$$

$$[\text{H}^+] = 0.016 \text{ M}$$

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



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

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

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