1.11.5: Water - Acid and Base in One

Learning Objectives

- Describe the autoionization of water.
- Calculate the concentrations of H_3O^+ and OH^- in aqueous solutions, knowing the other concentration.

We have already seen that H_2O can act as an acid or a base:

$$\underbrace{NH_3}_{\text{base}} + \underbrace{H_2O}_{\text{acid}} \rightleftharpoons NH_4^+ + OH^-$$

where H_2O acts as an acid (in red).

$$\underbrace{HCl}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \longrightarrow H_3O^+ + Cl^-$$

where H_2O acts as an base (in blue).

It may not surprise you to learn, then, that within any given sample of water, some H_2O molecules are acting as acids, and other H_2O molecules are acting as bases. The chemical equation is as follows:

$$\underbrace{H_2O}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \rightleftharpoons H_3O^+ + OH^- \tag{1.11.5.1}$$

This occurs only to a very small degree: only about 6 in 10^8 H_2O molecules are participating in this process, which is called the **autoionization of water**.

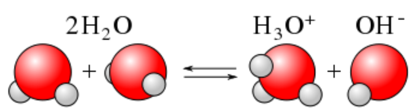

Figure 1.11.5.1: Autoionization of water, resulting in hydroxide and hydronium ions.

At this level, the concentration of both $H_3O^+(aq)$ and $OH^-(aq)$ in a sample of pure H_2O is about $1.0 \times 10^{-7}\ M$ (at room temperature). If we use square brackets—[ ]—around a dissolved species to imply the molar concentration of that species, we have

$$[H_3O^+] = [OH^-] = 1.0 \times 10^{-7} \tag{1.11.5.2}$$

for *any* sample of pure water because H2O can act as both an acid and a base. The product of these two concentrations is 1.0×10^{-14}:

$$[H_3O^+] \times [OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

- For acids, the concentration of $H_3O^+(aq)$ (i.e., $[H_3O^+]$) is greater than $1.0 \times 10^{-7}\ M$.
- For bases the concentration of $OH^-(aq)$ (i.e., $[OH^-]$) is greater than $1.0 \times 10^{-7}\ M$.

However, the *product* of the two concentrations—$[H_3O^+][OH^-]$—is *always* equal to 1.0×10^{-14}, no matter whether the aqueous solution is an acid, a base, or neutral:

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

This value of the product of concentrations is so important for aqueous solutions that it is called the **autoionization constant of water** and is denoted K_w:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \tag{1.11.5.3}$$

This means that if you know $[H_3O^+]$ for a solution, you can calculate what $[OH^-]$) has to be for the product to equal 1.0×10^{-14}; or if you know $[OH^-]$), you can calculate $[H_3O^+]$. This also implies that as one concentration goes up, the other must go down to compensate so that their product always equals the value of K_w.

Warning: Temperature Matters

The degree of autoionization of water (Equation 1.11.5.1)—and hence the value of K_w—changes with temperature, so Equations 1.11.5.2- 1.11.5.3are accurate only at room temperature.

Example 1.11.5.1: Hydroxide Concentration

What is $[OH^-]$) of an aqueous solution if $[H_3O^+]$ is $1.0 \times 10^{-4}\, M$?

Solution

Solutions to Example 14.7.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $[H_3O^+] = 1.0 \times 10^{-4}\ M$ Find: $[OH^-]$ = ? M
List other known quantities.	none
Plan the problem.	Using the expression for K_w, (Equation 1.11.5.3), rearrange the equation algebraically to solve for $[OH^-]$. $$[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]}$$
Calculate.	Now substitute the known quantities into the equation and solve. $$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = 1.0 \times 10^{-10}\, M$$ It is assumed that the concentration unit is molarity, so $[OH^-]$ is 1.0 × 10^{-10} M.
Think about your result.	The concentration of the acid is high (> 1 x 10^{-7} M), so $[OH^-]$ should be low.

Exercise 1.11.5.1

What is $[OH^-]$ in a 0.00032 M solution of H_2SO_4?

Hint

Assume **both** protons ionize from the molecule...although this is not the case.

Answer

$3.1 \times 10^{-11}\, M$

When you have a solution of a particular acid or base, you need to look at the formula of the acid or base to determine the number of H_3O^+ or OH^- ions in the formula unit because $[H_3O^+]$ or $[OH^-]$) may not be the same as the concentration of the acid or base itself.

Example 1.11.5.2: Hydronium Concentration

What is $[H_3O^+]$ in a 0.0044 M solution of $Ca(OH)_2$?

Solution

Solutions to Example 14.7.2

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $[Ca(OH)_2] = 0.0044\ M$ Find: $[H_3O^+]$ = ? M

Steps for Problem Solving	
List other known quantities.	We begin by determining $[\mathrm{OH^-}]$. The concentration of the solute is 0.0044 M, but because $\mathrm{Ca(OH)_2}$ is a strong base, there are two OH^- ions in solution for every formula unit dissolved, so the actual $[\mathrm{OH^-}]$ is two times this: $$[\mathrm{OH^-}]=2 \times 0 \cdot 0044 \mathrm{~M}=0 \cdot 0088 \mathrm{~M} \cdot$$
Plan the problem.	Use the expression for K_w (Equation 1.11.5.3) and rearrange the equation algebraically to solve for $[\mathrm{H_3O^+}]$. $$\left[H_3O^+\right] = \frac{1.0 \times 10^{-14}}{\left[OH^-\right]}$$
Calculate.	Now substitute the known quantities into the equation and solve. $$\left[H_3O^+\right] = \frac{1.0 \times 10^{-14}}{(0.0088)} = 1.1 \times 10^{-12} M$$ $[\mathrm{H_3O^+}]$ has decreased significantly in this basic solution.
Think about your result.	The concentration of the base is high (> 1 x 10^{-7} M) so $[\mathrm{H_3O^+}]$ should be low.

Exercise 1.11.5.2

What is $[\mathrm{H_3O^+}]$ of an aqueous solution if $[\mathrm{OH^-}]$ is $1.0 \times 10^{-9}\ M$?

Answer

1.0×10^{-5} M

In any aqueous solution, the product of $[\mathrm{H_3O^+}]$ and $[\mathrm{OH^-}]$ equals 1.0×10^{-14} (at room temperature).

Contributions & Attributions