

15.6: Water as an Acid and as a Base

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**, as with its reaction with NH_3 :



and that H_2O can act as a **base**, as with its reaction with HCl :



Therefore, it may not surprise you to learn that within any given sample of water, some H_2O molecules act as an **acid** and some H_2O molecules act as a **base**. The chemical equation is as follows:



However, this process (shown in [Figure 15.6.1](#)), called the **autoionization of water**, occurs to a very small extent, with only 6 out of every 100,000,000 H_2O molecules participating in autoionization at 25°C .

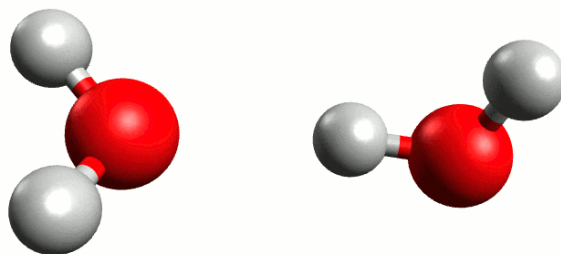


Figure 15.6.1: The autoionization of water, resulting in hydronium ions and hydroxide ions. (Manuel Almagro Rivas via Wikimedia Commons)

This represents a concentration of $1.0 \times 10^{-7} \text{ M}$ for both $\text{H}_3\text{O}^+ (\text{aq})$ and $\text{OH}^- (\text{aq})$ in a sample of pure H_2O (at 25°C). As [previously discussed](#), square brackets, $[]$, around a dissolved species represents the molarity of that species. This means that for any sample of pure water:

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M} \quad (15.6.4)$$

The product of these two concentrations is 1.0×10^{-14} :

$$[\text{H}_3\text{O}^+] \times [\text{OH}^-] = (1.0 \times 10^{-7}) (1.0 \times 10^{-7}) = 1.0 \times 10^{-14} \quad (15.6.5)$$

It turns out that for an aqueous solution at 25°C , the *product* of the two concentrations, $[\text{H}_3\text{O}^+] \times [\text{OH}^-]$, is *always* equal to 1.0×10^{-14} , regardless of whether the aqueous solution is acidic, basic, or neutral. In other words,

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} \quad (15.6.6)$$

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

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} \quad (15.6.7)$$

This means that if $[\text{H}_3\text{O}^+]$ is known for an aqueous solution, $[\text{OH}^-]$ may be calculated, since their product is always 1.0×10^{-14} ; or if $[\text{OH}^-]$ is known, that $[\text{H}_3\text{O}^+]$ may be calculated. This also implies that as one concentration goes up, the other must go down to ensure their product always equals the value of K_w .

Temperature Matters

The degree of autoionization of water, as shown in [Equation 15.6.3](#) changes with temperature. This means that the value of K_w also changes with temperature. Therefore, [Equation 15.6.4](#) through [Equation 15.6.7](#) are only valid at room temperature (25°C). In the event the temperature is not known, it is assumed to be 25°C.

Since pure water is neutral and $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$ in pure water, then any aqueous solution in which $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$ must also be neutral (at 25°C). From this, we can conclude the following:

- Since acids produce H_3O^+ when dissolved in water, $[\text{H}_3\text{O}^+] > 1.0 \times 10^{-7} \text{ M}$ in acidic solutions (at 25°C).
- Since bases produce OH^- when dissolved in water, $[\text{OH}^-] > 1.0 \times 10^{-7} \text{ M}$ in basic solutions (at 25°C).

Example 15.6.1

What is the $[\text{OH}^-]$ of an aqueous solution if the $[\text{H}_3\text{O}^+]$ is $1.0 \times 10^{-4} \text{ M}$?

Solution

Steps for Problem Solving

Identify the "given" information and what the problem is asking you to "find."

Given: $[\text{H}_3\text{O}^+] = 1.0 \times 10^{-4} \text{ M}$

Find: $[\text{OH}^-] = ? \text{ M}$

List known relationship(s).

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

Plan the problem.

Rearrange the equation to solve for $[\text{OH}^-]$.

$$[\text{OH}^-] = \frac{\cancel{[\text{H}_3\text{O}^+]}[\text{OH}^-]}{\cancel{[\text{H}_3\text{O}^+]}} = \frac{1.0 \times 10^{-14}}{[\text{H}_3\text{O}^+]}$$

Calculate the answer.

Now substitute the known quantities into the equation and solve.

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = \boxed{1.0 \times 10^{-10} \text{ M}}$$

Steps for Problem Solving

Think about your result.

The $[\text{H}_3\text{O}^+]$ is high ($> 1 \times 10^{-7} \text{ M}$), so $[\text{OH}^-]$ should be low.

Furthermore, a quick answer check shows:

$$[\text{H}_3\text{O}^+][\text{OH}^-] = (1.0 \times 10^{-4})(1.0 \times 10^{-10}) = 1.0 \times 10^{-14}$$

Exercise 15.6.1

What is $[\text{H}_3\text{O}^+]$ in solution where $[\text{OH}^-] = 0.00032 \text{ M}$?

Answer

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

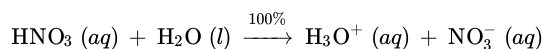
$[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ of Strong Acids and Bases

Finding the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ of *strong* acids and *strong* bases is relatively straightforward. Finding the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ of *weak* acids and *weak* bases is a bit more complicated and beyond the scope of this text.

As [previously discussed](#), strong acids ionize completely. For *monoprotic* strong acids (strong acids with only one H),

$$[\text{H}_3\text{O}^+] = [\text{acid}] \quad (15.6.8)$$

This means that $[\text{H}_3\text{O}^+] = 0.10 \text{ M}$ in a solution of 0.10 M HNO_3 , a strong acid, since one H_3O^+ ion forms for each molecule of HNO_3 :

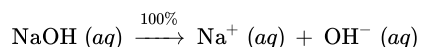


For H_2SO_4 , the only strong acid that is *diprotic* (has two H atoms), $[\text{H}_3\text{O}^+] = [\text{acid}]$ is a close enough approximation, though the $[\text{H}_3\text{O}^+]$ is slightly greater than $[\text{acid}]$, since more than one H_3O^+ ion forms for each molecule of H_2SO_4 . However, it is much more difficult for the second proton (H^+) to be released from H_2SO_4 , which is why $[\text{H}_3\text{O}^+] = [\text{acid}]$ remains a close enough approximation.

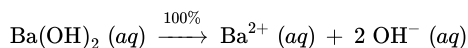
Also [previously discussed](#), strong bases are soluble in water and dissociate completely. For strong bases,

$$[\text{OH}^-] = [\text{base}] \times \text{number of OH}^- \text{ ions in the chemical formula} \quad (15.6.9)$$

This means that $[\text{OH}^-] = 0.10 \text{ M}$ in a solution of 0.10 M NaOH , a strong base, since there is only one OH^- ion in NaOH :



and that $[\text{OH}^-] = 0.20 \text{ M}$ in a solution of 0.10 M Ba(OH)_2 , also a strong base, since there are two OH^- ions in Ba(OH)_2 :



Example 15.6.2

What is the:

- A. $[\text{H}_3\text{O}^+]$ in 0.0023 M HBr ?
- B. $[\text{H}_3\text{O}^+]$ in 0.0023 M HNO_3 ?
- C. $[\text{OH}^-]$ in 0.0023 M KOH ?
- D. $[\text{OH}^-]$ in $0.0023 \text{ M Sr(OH)}_2$?

Solution

- A. $[\text{H}_3\text{O}^+] = [\text{acid}]$ in a strong acid, so $[\text{H}_3\text{O}^+] = 0.0023 \text{ M}$
- B. $[\text{H}_3\text{O}^+] = [\text{acid}]$ in a strong acid, so $[\text{H}_3\text{O}^+] = 0.0023 \text{ M}$
- C. $[\text{OH}^-] = [\text{base}] \times \text{number of OH}^- \text{ ions in the chemical formula}$ for a strong base, so $[\text{OH}^-] = 0.0023 \text{ M}$, since there is only one OH^- ion in KOH .
- D. $[\text{OH}^-] = [\text{base}] \times \text{number of OH}^- \text{ ions in the chemical formula}$ for a strong base, so $[\text{OH}^-] = 0.0046 \text{ M}$, since there are two OH^- ions in $\text{Sr}(\text{OH})_2$. ($0.0023 \text{ M} \times 2 = 0.0046 \text{ M}$).

✓ Example 15.6.3

What is $[\text{H}_3\text{O}^+]$ in a 0.0044 M solution of $\text{Ca}(\text{OH})_2$?

Solution

Steps for Problem Solving

| | |
|--|---|
| Identify the "given" information and what the problem is asking you to "find." | <p>Given: $\text{Ca}(\text{OH})_2 = 0.0044 \text{ M}$</p> <p>Find: $[\text{H}_3\text{O}^+] = ? \text{ M}$</p> |
| List known relationship(s). | <p>$[\text{OH}^-] = [\text{base}] \times \text{number of OH}^- \text{ ions in the chemical formula}$</p> <p>$[\text{H}_3\text{O}^+] [\text{OH}^-] = 1.0 \times 10^{-14}$</p> |
| Plan the problem. | <p>Rearrange the equation to solve for $[\text{H}_3\text{O}^+]$.</p> $[\text{H}_3\text{O}^+] = \frac{[\text{H}_3\text{O}^+] \cancel{[\text{OH}^-]}}{\cancel{[\text{OH}^-]}} = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]}$ |
| Calculate. | <p>$[\text{OH}^-] = 0.0044 \text{ M} \times 2 = 0.0088 \text{ M}$</p> $[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.0088} = \boxed{1.1 \times 10^{-12} \text{ M}}$ |
| Think about your result. | <p>The $[\text{OH}^-]$ is high ($> 1 \times 10^{-7} \text{ M}$), so $[\text{H}_3\text{O}^+]$ should be low ($< 1 \times 10^{-7} \text{ M}$).</p> |

✏ Exercise 15.6.3

Calculate the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ of $6.1 \times 10^{-3} \text{ M HI}$.

Answer

$$[\text{H}_3\text{O}^+] = 6.1 \times 10^{-3} \text{ M}; [\text{OH}^-] = 1.6 \times 10^{-12} \text{ M}$$

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