The pH of an aqueous solution is the measure of how acidic or basic it is. The pH of an aqueous solution can be determined and calculated by using the concentration of hydronium ion concentration in the solution.

**Introduction**

The pH of an aqueous solution is based on the pH scale which typically ranges from 0 to 14 in water (although as discussed below this is not an a formal rule). A pH of 7 is considered to be neutral. A pH of less than 7 is considered acidic. A pH of greater than 7 is then considered basic. Acidic solutions have high hydronium concentrations and lower hydroxide concentrations. Basic solutions have high hydroxide concentrations and lower hydronium concentrations.

**Self-Ionization of Water**

In the self-ionization of water, the amphiprotic ability of water to act as a proton donor and acceptor allows the formation of hydronium ($H_3O^+$) and hydroxide ions ($OH^-$). In pure water, the concentration of hydronium ions equals that of hydroxide ions. At 25 °C, the concentrations of both hydronium and hydroxide ions equal $1.0 \times 10^{-7}$. The ion product of water, $K_w$, is the equilibrium condition for the self-ionization of water and is express as follows:

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

- **pH**: The term pH refers to the "potential of hydrogen ion." It was proposed by Danish biochemist Soren Sorensen in 1909 so that there could be a more convenient way to describe hydronium and hydroxide ion concentrations in aqueous solutions since both concentrations tend to be extremely small. Sorensen defined pH as the negative of the logarithm of the concentration of hydrogen ions. In terms of hydronium ion concentration, the equation to determine the pH of an aqueous solution is: $\text{pH} = -\log[H_3O^+]$

- **pOH**: The pOH of an aqueous solution, which is related to the pH, can be determined by the following equation: $\text{pOH} = -\log[OH^-]$ This equation uses the hydroxide concentration of an aqueous solution instead of the hydronium concentration.

**Relating pH and pOH**

Another equation can be used that relates the concentrations of hydronium and hydroxide concentrations. This equation is derived from the equilibrium condition for the self-ionization of water, $K_w$. It brings the three equations for pH, pOH, and $K_w$ together to show that they are all related to each other and either one can be found if the other two are known. The following equation is expressed by taking the negative logarithm of the $K_w$ expression for the self-ionization of water at room temperature:

$$\text{pK}_w = \text{pH} + \text{pOH} = 14$$
Strong Acids and Strong Bases

The ionization of strong acids and strong bases in dilute aqueous solutions essentially go to completion. In aqueous solutions of strong acids and strong bases, the self-ionization of water only occurs to a small extent. Since it only occurs to a small extent, the self-ionization of water is an insignificant source of hydronium and hydroxide ions. Knowing this, we can say in calculating hydronium concentration in an aqueous solution of a strong acid that the strong acid is the main source of hydronium ions. We can also say that in calculating hydroxide concentration in an aqueous solution of a strong base that the strong base is the main source of hydroxide ions. This is usually true unless the solutions of strong acids and strong bases are extremely dilute.

Weak Acids and Weak Bases

Weak acids only partially dissociate in aqueous solutions and reach a condition of equilibrium, therefore how much they dissociate is given by the equilibrium equation for that acid in solution:

\[ K_a = \dfrac{[H_3O^+][A^-]}{[HA]} \]

with

- \([H_3O^+]\) is the Hydronium Concentration
- \([A^-]\) is conjugate base concentration
- \([HA]\) is the Weak Acid concentration

Weak bases also only partially dissociate in aqueous solutions and reach a condition of equilibrium. The equation for the partial dissociation of a base is then the equilibrium equation for that base in solution:

\[ K_b = \dfrac{[OH^-][B^+]}{[B]} \]

\([OH^-]\) = Hydroxide Concentration

\([B^+]=\text{ion}\]

\([B]=\text{Weak Base}\]

References


Problems

1. A solution is 0.055 M HBr. What is the pH of this solution?

2. A solution is 0.025 M NaOH. What is the pH of this solution?
2. A solution is 0.00025 M HCl. What is the pH AND pOH of this solution?

3. A solution is 0.0035 M LiOH. What is the pOH of this solution? pH?

4. A solution contains 0.0045 M hydrofluoric acid. What is the pH of this solution? For hydrofluoric acid, \(K_a = 6.6 \times 10^{-4}\).

5. A solution contains 0.0085 M ammonia. What is the pH of this solution? For ammonia: \(K_b = 1.8 \times 10^{-5}\).

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### Answers

1. Use the pH equation which is: \(pH = -\log[H_3O^+]\).

   0.055 M HBr, HBr is a strong acid

   \[ [H_3O^+] = 5.5 \times 10^{-2} \text{ M} \]

   \[ pH = -\log(5.5 \times 10^{-2}) = 1.26 \]

2. Use the pH equation \(pH = -\log[H_3O^+]\) and pK\(_w\) equation \(pK_w = pH + pOH = 14\).

   0.00025 M HCl, HCl is a strong acid

   \[ [H_3O^+] = 2.5 \times 10^{-4} \text{ M} \]

   \[ pH = -\log(2.5 \times 10^{-4}) = 3.6 \]

   Then solve for the pOH:

   \[ pH + pOH = 14 \]

   \[ pOH = 14 - pH \]

   \[ pOH = 14 - 3.6 = 10.4 \]

3. Use the pOH equation \(pH = -\log[OH^-]\) and pK\(_w\) equation \(pK_w = pH + pOH = 14\).

   0.0035 M LiOH, LiOH is a strong base

   \[ [OH^-] = 3.5 \times 10^{-3} \]

   \[ pOH = -\log(3.5 \times 10^{-3}) = 2.46 \]

   Now solve for pH:

   \[ pH = 14 - pOH \]

   \[ pH = 14 - 2.46 = 11.54 \]
4. 0.0045 M hydrofluoric acid, hydrofluoric acid is a weak acid.

Use $K_a$ equation \( K_a = \frac{[H_3O^+][A^-]}{[HA]} \) and ICE table.

5. 0.0085 M ammonia, ammonia is a weak base.

Use $K_b$ equation \( K_b = \frac{[OH^-][B^+]}{[B]} \) and ICE table.

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