

Dissociation of Water

- Occasionally, a hydrogen atom shared by two water molecules shifts from one molecule to the other.
 - The hydrogen atom leaves its electron behind and is transferred as a single proton - a **hydrogen ion (H⁺)**.
 - When the hydrogen ion (H⁺) bonds with another water molecule, it forms a **hydronium ion (H₃O⁺)**
 - The water molecule that lost a proton is now a **hydroxide ion (OH⁻)**.
 - $\text{H}_2\text{O} \rightarrow \text{H}^+ + \text{OH}^-$
 - $\text{H}^+ + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+$
- Another way to view this process is that a water molecule dissociates into a hydrogen ion and a hydroxide ion:
 - $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$
- This reaction is reversible.
- At equilibrium the concentration of water molecules (H₂O) greatly exceeds that of H⁺ and OH⁻.
- In *pure water* only one water molecule in every 554 million is dissociated.
- Because hydrogen and hydroxide ions are very reactive, changes in their concentrations can drastically affect the proteins and other molecules of a cell.
- Adding certain solutes, called acids and bases, disrupts the equilibrium and modifies the concentrations of hydrogen and hydroxide ions.
- The pH scale is used to describe how acidic or basic (the opposite of acidic) a solution is.

Organisms are sensitive to changes in pH

- An **acid** is a substance that increases the hydrogen ion (H⁺) concentration in a solution.
- Addition of an acid makes a solution more acidic.
- Any substance that reduces the hydrogen ion (H⁺) concentration in a solution is a **base**.
- Solutions with more OH⁻ than H⁺ are basic solutions
- Some acids and bases (HCl and NaOH) are strong acids or bases.
 - These molecules dissociate completely in water.
- Other acids and bases are weak acids or bases.
 - For these molecules, the binding and release of hydrogen ions are reversible.
 - At equilibrium there will be a fixed ratio of products to reactants.
 - Carbonic acid (H₂CO₃) is a weak acid:
 - $\text{H}_2\text{CO}_3 \rightleftharpoons \text{HCO}_3^- + \text{H}^+$
 - At equilibrium, 1% of the molecules will be dissociated.
- In any solution the *product* of their H⁺ and OH⁻ concentrations is constant at 10⁻¹⁴M.

Chapter 3-2: Dissociation and Ph

- $[H^+][OH^-] = 10^{-14}$
 - In a neutral solution, $[H^+] = 10^{-7} M$ and $[OH^-] = 10^{-7} M$
 - $10^{-7} M = 1/10,000,000$ moles of H^+ ions (pH 7)
- Adding acid to a solution shifts the balance between H^+ and OH^- toward H^+ and leads to a decline in OH^- .
 - $[H^+] > [OH^-]$ for an acid
 - If $[H^+] = 10^{-5} M$ (pH 5), then $[OH^-] = 10^{-9} M$ (pH 9)
- Adding a base does the opposite, increasing OH^- concentration and dropping H^+ concentration.
 - $[H^+] < [OH^-]$ for a base
- The H^+ and OH^- concentrations of solutions can vary by a factor of 100 trillion or more.
- To express this variation more conveniently, the H^+ and OH^- concentrations are typically expressed via the **pH scale**.
 - The pH scale, ranging from 1 to 14, compresses the range of concentrations by employing logarithms.
 - $pH = -\log [H^+]$ or $[H^+] = 10^{-pH}$
- The pH of a neutral solution is 7.
- Acidic solutions have pH values less than 7 and basic solutions have pH values more than 7.
- Most biological fluids have pH values in the range of 6-8.
- Each pH unit represents a tenfold difference in H^+ and OH^- concentrations.
 - A small change in pH actually indicates a substantial change in H^+ and OH^- concentrations.
- The chemical processes in the cell can be disrupted by changes to the H^+ and OH^- concentrations away from their normal values near pH 7.
- To maintain cellular pH values at a constant level, biological fluids have buffers.
- **Buffers** resist changes to the pH of a solution when H^+ or OH^- is added to the solution.
 - Buffers accept hydrogen ions from the solution when they are in excess and donate hydrogen ions when they have been depleted.
- Buffers typically consist of a weak acid and its corresponding base.