

Chemical Equilibrium

Chemical equilibrium is a state in which the concentrations of reactants and products remain constant over time. It is a dynamic equilibrium, meaning that the reactions are still occurring, but the rates of the forward and reverse reactions are equal.



At equilibrium, the concentrations of reactants and products are related by the equilibrium constant, K .

The equilibrium constant, K , is a measure of the extent to which a reaction proceeds. It is defined as the ratio of the concentrations of the products to the concentrations of the reactants, each raised to the power of its stoichiometric coefficient.

$$K = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

For a given reaction, K is constant at a given temperature. If $K > 1$, the reaction favors the products. If $K < 1$, the reaction favors the reactants.

The equilibrium constant can be used to predict the direction in which a reaction will proceed. If the reaction quotient, Q , is less than K , the reaction will proceed in the forward direction. If $Q > K$, the reaction will proceed in the reverse direction.



The equilibrium constant can be used to calculate the concentrations of reactants and products at equilibrium. For example, if the equilibrium constant for a reaction is $K = 10$, and the initial concentrations of reactants are $[A] = 1.0 \text{ M}$ and $[B] = 1.0 \text{ M}$, the equilibrium concentrations of reactants and products can be calculated.



The equilibrium constant can be used to calculate the standard Gibbs free energy change, ΔG° , for a reaction. The relationship between K and ΔG° is given by the equation:

$$\Delta G^\circ = -RT \ln K$$

Le Chatelier's Principle

Le Chatelier's Principle states that if a system at equilibrium is subjected to a change in concentration, pressure, or temperature, the system will shift in a direction that opposes the change and re-establishes equilibrium.

For example, if the concentration of a reactant is increased, the system will shift in the forward direction to consume the excess reactant and produce more products.

Similarly, if the pressure is increased, the system will shift in the direction that produces fewer moles of gas.

Finally, if the temperature is increased, the system will shift in the direction that absorbs heat.

Le Chatelier's Principle can be used to predict the effect of changes in concentration, pressure, and temperature on the equilibrium position of a reaction.

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Acid-Base Equilibrium

Acid-base equilibrium is a type of chemical equilibrium that involves the transfer of protons between acids and bases. The equilibrium constant for an acid-base reaction is called the acid dissociation constant, K_a .

The acid dissociation constant, K_a , is a measure of the strength of an acid. It is defined as the ratio of the concentrations of the products to the concentrations of the reactants, each raised to the power of its stoichiometric coefficient.



For a given acid, K_a is constant at a given temperature. If $K_a > 1$, the acid is strong. If $K_a < 1$, the acid is weak.

The acid dissociation constant can be used to predict the pH of a solution. For example, if the K_a for an acid is 1.0×10^{-4} , and the initial concentration of the acid is 1.0 M , the pH of the solution can be calculated.



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Buffer Solutions

A buffer solution is a solution that resists changes in pH when small amounts of acid or base are added. Buffer solutions are typically composed of a weak acid and its conjugate base, or a weak base and its conjugate acid.

The pH of a buffer solution can be calculated using the Henderson-Hasselbalch equation:

$$\text{pH} = \text{p}K_a + \log \frac{[A^-]}{[HA]}$$

For example, if the $\text{p}K_a$ for an acid is 4.0, and the initial concentrations of the acid and its conjugate base are $[HA] = 1.0 \text{ M}$ and $[A^-] = 1.0 \text{ M}$, the pH of the buffer solution can be calculated.



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