



CHEMISTRY

CHEAT SHEET

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Chemical Equilibrium occurs when the concentrations of the chemical species remain the same with time (dynamic equilibrium). Equilibrium reactions are represented by the symbol \rightleftharpoons .

<p>What is Equilibrium?</p> <p>Equilibrium is a state of balance between two or more opposing processes. In chemistry, it refers to a state where the concentrations of reactants and products remain constant over time.</p> <p>Equilibrium is a dynamic process, meaning that the reactions are still occurring, but the rates of the forward and reverse reactions are equal.</p> <p>Equilibrium is reached when the system has no net change in the concentrations of the reactants and products.</p>	<p>Equilibrium Constant</p> <p>The equilibrium constant (K) is a measure of the extent to which a reaction proceeds towards equilibrium. 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.</p> <p>For a general reaction:</p> $aA + bB \rightleftharpoons cC + dD$ <p>The equilibrium constant is given by:</p> $K = \frac{[C]^c [D]^d}{[A]^a [B]^b}$	<p>Graphical Representation of Equilibrium</p>												
<p>Equilibrium Constant, K_c</p> <p>The equilibrium constant K_c is calculated using the concentrations of the reactants and products in moles per liter (M).</p> <p>For the reaction:</p> $2H_2(g) + O_2(g) \rightleftharpoons 2H_2O(g)$ <p>The equilibrium constant K_c is given by:</p> $K_c = \frac{[H_2O]^2}{[H_2]^2 [O_2]}$	<p>Equilibrium Constant, K_p</p> <p>The equilibrium constant K_p is calculated using the partial pressures of the reactants and products in atmospheres (atm).</p> <p>For the reaction:</p> $2H_2(g) + O_2(g) \rightleftharpoons 2H_2O(g)$ <p>The equilibrium constant K_p is given by:</p> $K_p = \frac{(P_{H_2O})^2}{(P_{H_2})^2 (P_{O_2})}$	<p>Equilibrium Constant, K_c vs. K_p</p> <p>The relationship between K_c and K_p is given by the equation:</p> $K_p = K_c (RT)^{\Delta n}$ <p>where R is the gas constant, T is the temperature in Kelvin, and Δn is the change in the number of moles of gas.</p>												
<p>Equilibrium Constant, K_a</p> <p>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 for the dissociation of an acid.</p> <p>For the dissociation of an acid:</p> $HA \rightleftharpoons H^+ + A^-$ <p>The acid dissociation constant K_a is given by:</p> $K_a = \frac{[H^+] [A^-]}{[HA]}$	<p>Equilibrium Constant, K_b</p> <p>The base dissociation constant K_b is a measure of the strength of a base. It is defined as the ratio of the concentrations of the products to the concentrations of the reactants for the dissociation of a base.</p> <p>For the dissociation of a base:</p> $B + H_2O \rightleftharpoons BH^+ + OH^-$ <p>The base dissociation constant K_b is given by:</p> $K_b = \frac{[BH^+] [OH^-]}{[B] [H_2O]}$	<p>Equilibrium Constant, K_a vs. K_b</p> <p>The relationship between K_a and K_b for a conjugate acid-base pair is given by the equation:</p> $K_a \times K_b = K_w$ <p>where K_w is the ion-product constant for water, which is 1.0×10^{-14} at 25°C.</p>												
<p>Le Chatelier's Principle: the effect of changing conditions on equilibrium</p> <table border="1"> <thead> <tr> <th>Condition</th> <th>Effect on Equilibrium</th> <th>Example</th> </tr> </thead> <tbody> <tr> <td>Concentration</td> <td> <ul style="list-style-type: none"> Increasing the concentration of a reactant shifts the equilibrium to the right (towards the products). Increasing the concentration of a product shifts the equilibrium to the left (towards the reactants). </td> <td> <p>For the reaction:</p> $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ <p>Increasing the concentration of N_2 or H_2 will shift the equilibrium to the right, increasing the concentration of NH_3.</p> </td> </tr> <tr> <td>Pressure</td> <td> <ul style="list-style-type: none"> Increasing the pressure shifts the equilibrium towards the side with fewer moles of gas. Decreasing the pressure shifts the equilibrium towards the side with more moles of gas. </td> <td> <p>For the reaction:</p> $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ <p>Increasing the pressure will shift the equilibrium to the right, as there are fewer moles of gas on the product side (2 moles) compared to the reactant side (4 moles).</p> </td> </tr> <tr> <td>Temperature</td> <td> <ul style="list-style-type: none"> Increasing the temperature shifts the equilibrium towards the endothermic direction. Decreasing the temperature shifts the equilibrium towards the exothermic direction. </td> <td> <p>For the reaction:</p> $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ <p>The forward reaction is exothermic. Increasing the temperature will shift the equilibrium to the left, decreasing the concentration of NH_3.</p> </td> </tr> </tbody> </table>			Condition	Effect on Equilibrium	Example	Concentration	<ul style="list-style-type: none"> Increasing the concentration of a reactant shifts the equilibrium to the right (towards the products). Increasing the concentration of a product shifts the equilibrium to the left (towards the reactants). 	<p>For the reaction:</p> $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ <p>Increasing the concentration of N_2 or H_2 will shift the equilibrium to the right, increasing the concentration of NH_3.</p>	Pressure	<ul style="list-style-type: none"> Increasing the pressure shifts the equilibrium towards the side with fewer moles of gas. Decreasing the pressure shifts the equilibrium towards the side with more moles of gas. 	<p>For the reaction:</p> $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ <p>Increasing the pressure will shift the equilibrium to the right, as there are fewer moles of gas on the product side (2 moles) compared to the reactant side (4 moles).</p>	Temperature	<ul style="list-style-type: none"> Increasing the temperature shifts the equilibrium towards the endothermic direction. Decreasing the temperature shifts the equilibrium towards the exothermic direction. 	<p>For the reaction:</p> $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ <p>The forward reaction is exothermic. Increasing the temperature will shift the equilibrium to the left, decreasing the concentration of NH_3.</p>
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Chemistry Chemical Kinetics And Equilibrium Study Guide

Chao Zhang



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