



# CHEMISTRY

## CHEAT SHEET

# EQUILIBRIUM

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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><b>What is Equilibrium?</b></p> <p>Equilibrium is a state of balance between two or more opposing processes. In a chemical reaction, it is the state where the concentrations of the reactants and products remain constant over time.</p> <p>Equilibrium is a dynamic process, meaning that the reaction is still occurring, but the rates of the forward and reverse reactions are equal.</p>	<p><b>Equilibrium Constant</b></p> <p>The equilibrium constant, <math>K</math>, 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.</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><b>Graphical Representation of Equilibrium</b></p>												
<p><b>Equilibrium Constant, <math>K_c</math></b></p> <p><math>K_c</math> is the equilibrium constant in terms of concentration. It is calculated using the concentrations of the reactants and products at equilibrium.</p> <p>For a general reaction:</p> $aA + bB \rightleftharpoons cC + dD$ <p>The equilibrium constant is given by:</p> $K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$	<p><b>Equilibrium Constant, <math>K_p</math></b></p> <p><math>K_p</math> is the equilibrium constant in terms of partial pressure. It is calculated using the partial pressures of the reactants and products at equilibrium.</p> <p>For a general reaction:</p> $aA + bB \rightleftharpoons cC + dD$ <p>The equilibrium constant is given by:</p> $K_p = \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b}$	<p><b>Equilibrium Constant, <math>K_c</math> vs. <math>K_p</math></b></p> <p>The relationship between <math>K_c</math> and <math>K_p</math> is given by the equation:</p> $K_p = K_c (RT)^{\Delta n}$ <p>where <math>R</math> is the gas constant, <math>T</math> is the temperature in Kelvin, and <math>\Delta n</math> is the change in the number of moles of gas.</p>												
<p><b>Equilibrium Constant, <math>K_a</math></b></p> <p><math>K_a</math> is the acid dissociation constant. It is a measure of the strength of an acid in solution. It is calculated using the concentrations of the acid, its conjugate base, and the hydrogen ions at equilibrium.</p> <p>For a general acid dissociation reaction:</p> $HA \rightleftharpoons H^+ + A^-$ <p>The equilibrium constant is given by:</p> $K_a = \frac{[H^+] [A^-]}{[HA]}$	<p><b>Equilibrium Constant, <math>K_b</math></b></p> <p><math>K_b</math> is the base dissociation constant. It is a measure of the strength of a base in solution. It is calculated using the concentrations of the base, its conjugate acid, and the hydroxide ions at equilibrium.</p> <p>For a general base dissociation reaction:</p> $B + H_2O \rightleftharpoons BH^+ + OH^-$ <p>The equilibrium constant is given by:</p> $K_b = \frac{[BH^+] [OH^-]}{[B]}$	<p><b>Equilibrium Constant, <math>K_a</math> vs. <math>K_b</math></b></p> <p>The relationship between <math>K_a</math> and <math>K_b</math> is given by the equation:</p> $K_a \cdot K_b = K_w$ <p>where <math>K_w</math> is the ion-product constant for water, which is equal to <math>1.0 \times 10^{-14}</math> at 25°C.</p>												
<p><b>Le Chatelier's Principle: the effect of changing conditions on equilibrium</b></p> <table> <tr> <th>Condition</th><th>Effect on Equilibrium</th><th>Example</th></tr> <tr> <td>Concentration</td><td> <ul style="list-style-type: none"> <li>Increasing the concentration of a reactant will shift the equilibrium to the right, favoring the products.</li> <li>Increasing the concentration of a product will shift the equilibrium to the left, favoring the reactants.</li> </ul> </td><td> <p>For the reaction:</p> <math display="block">N_2 + 3H_2 \rightleftharpoons 2NH_3</math> <p>Increasing the concentration of <math>N_2</math> will shift the equilibrium to the right, producing more <math>NH_3</math>.</p> </td></tr> <tr> <td>Pressure</td><td> <ul style="list-style-type: none"> <li>Increasing the pressure will shift the equilibrium to the side with the fewest moles of gas.</li> <li>Decreasing the pressure will shift the equilibrium to the side with the most moles of gas.</li> </ul> </td><td> <p>For the reaction:</p> <math display="block">N_2 + 3H_2 \rightleftharpoons 2NH_3</math> <p>Increasing the pressure will shift the equilibrium to the right, producing more <math>NH_3</math>.</p> </td></tr> <tr> <td>Temperature</td><td> <ul style="list-style-type: none"> <li>Increasing the temperature will shift the equilibrium to the side that is endothermic (absorbs heat).</li> <li>Decreasing the temperature will shift the equilibrium to the side that is exothermic (releases heat).</li> </ul> </td><td> <p>For the reaction:</p> <math display="block">N_2 + 3H_2 \rightleftharpoons 2NH_3</math> <p>Increasing the temperature will shift the equilibrium to the left, favoring the reactants.</p> </td></tr> </table>			Condition	Effect on Equilibrium	Example	Concentration	<ul style="list-style-type: none"> <li>Increasing the concentration of a reactant will shift the equilibrium to the right, favoring the products.</li> <li>Increasing the concentration of a product will shift the equilibrium to the left, favoring the reactants.</li> </ul>	<p>For the reaction:</p> $N_2 + 3H_2 \rightleftharpoons 2NH_3$ <p>Increasing the concentration of <math>N_2</math> will shift the equilibrium to the right, producing more <math>NH_3</math>.</p>	Pressure	<ul style="list-style-type: none"> <li>Increasing the pressure will shift the equilibrium to the side with the fewest moles of gas.</li> <li>Decreasing the pressure will shift the equilibrium to the side with the most moles of gas.</li> </ul>	<p>For the reaction:</p> $N_2 + 3H_2 \rightleftharpoons 2NH_3$ <p>Increasing the pressure will shift the equilibrium to the right, producing more <math>NH_3</math>.</p>	Temperature	<ul style="list-style-type: none"> <li>Increasing the temperature will shift the equilibrium to the side that is endothermic (absorbs heat).</li> <li>Decreasing the temperature will shift the equilibrium to the side that is exothermic (releases heat).</li> </ul>	<p>For the reaction:</p> $N_2 + 3H_2 \rightleftharpoons 2NH_3$ <p>Increasing the temperature will shift the equilibrium to the left, favoring the reactants.</p>
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