



# CHEMISTRY

## CHEAT SHEET

# EQUILIBRIUM

## EQUILIBRIUM CHEAT SHEET

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.</p> <p>In chemistry, it refers to a reaction where the forward and reverse rates are equal, and the concentrations of reactants and products remain constant over time.</p> <p>Equilibrium is a dynamic process, meaning that the reaction is still occurring, but the net change is zero.</p>	<p><b>Equilibrium Constants</b></p> <p><math>K_c</math> (Concentration): <math>K_c = \frac{[Products]}{[Reactants]}</math></p> <p><math>K_p</math> (Pressure): <math>K_p = \frac{P_{Products}}{P_{Reactants}}</math></p> <p><math>K_a</math> (Acid Dissociation): <math>K_a = \frac{[H^+][A^-]}{[HA]}</math></p> <p><math>K_b</math> (Base Dissociation): <math>K_b = \frac{[OH^-][BH^+]}{[B]}</math></p>													
<p><b>Equilibrium Constant, <math>K_c</math></b></p> <p>For a general reaction: <math>aA + bB \rightleftharpoons cC + dD</math></p> <p>The equilibrium constant <math>K_c</math> is given by:</p> $K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$ <p>Where <math>[A]</math>, <math>[B]</math>, <math>[C]</math>, and <math>[D]</math> are the molar concentrations of the reactants and products at equilibrium.</p>	<p><b>Equilibrium Constant, <math>K_p</math></b></p> <p>For a general reaction: <math>aA + bB \rightleftharpoons cC + dD</math></p> <p>The equilibrium constant <math>K_p</math> is given by:</p> $K_p = \frac{P_C^c P_D^d}{P_A^a P_B^b}$ <p>Where <math>P_A</math>, <math>P_B</math>, <math>P_C</math>, and <math>P_D</math> are the partial pressures of the reactants and products at equilibrium.</p>													
<p><b>Equilibrium Constant, <math>K_a</math></b></p> <p>For a weak acid <math>HA</math> dissociating in water:</p> $HA \rightleftharpoons H^+ + A^-$ <p>The acid dissociation constant <math>K_a</math> is given by:</p> $K_a = \frac{[H^+][A^-]}{[HA]}$	<p><b>Equilibrium Constant, <math>K_b</math></b></p> <p>For a weak base <math>B</math> reacting with water:</p> $B + H_2O \rightleftharpoons BH^+ + OH^-$ <p>The base dissociation constant <math>K_b</math> is given by:</p> $K_b = \frac{[BH^+][OH^-]}{[B]}$													
<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 concentration of a reactant shifts the equilibrium to the right (towards products).</li><li>Increasing concentration of a product shifts the equilibrium to the left (towards reactants).</li></ul></td><td><math display="block">N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)</math><p>Adding more <math>N_2</math> shifts 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 pressure shifts the equilibrium towards the side with fewer moles of gas.</li><li>Decreasing pressure shifts the equilibrium towards the side with more moles of gas.</li></ul></td><td><math display="block">N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)</math><p>Increasing pressure shifts the equilibrium to the right, as there are fewer moles of gas on the product side.</p></td></tr><tr><td>Temperature</td><td><ul style="list-style-type: none"><li>For an exothermic reaction, increasing temperature shifts the equilibrium to the left.</li><li>For an endothermic reaction, increasing temperature shifts the equilibrium to the right.</li></ul></td><td><math display="block">N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92 \text{ kJ/mol}</math><p>Increasing temperature shifts the equilibrium to the left, as the reaction is exothermic.</p></td></tr></table>			Condition	Effect on Equilibrium	Example	Concentration	<ul style="list-style-type: none"><li>Increasing concentration of a reactant shifts the equilibrium to the right (towards products).</li><li>Increasing concentration of a product shifts the equilibrium to the left (towards reactants).</li></ul>	$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ <p>Adding more <math>N_2</math> shifts the equilibrium to the right, producing more <math>NH_3</math>.</p>	Pressure	<ul style="list-style-type: none"><li>Increasing pressure shifts the equilibrium towards the side with fewer moles of gas.</li><li>Decreasing pressure shifts the equilibrium towards the side with more moles of gas.</li></ul>	$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ <p>Increasing pressure shifts the equilibrium to the right, as there are fewer moles of gas on the product side.</p>	Temperature	<ul style="list-style-type: none"><li>For an exothermic reaction, increasing temperature shifts the equilibrium to the left.</li><li>For an endothermic reaction, increasing temperature shifts the equilibrium to the right.</li></ul>	$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92 \text{ kJ/mol}$ <p>Increasing temperature shifts the equilibrium to the left, as the reaction is exothermic.</p>
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# Chemistry Study Guide For Equilibrium And Thermochemistry

**Barron's Educational Series, Mark  
Kernion, Joseph A. Mascetta**



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