



# 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. 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 expressed in terms of molar concentrations. It is used for reactions in the gas phase and in aqueous solution.</p> <p>For a general reaction:</p> $aA + bB \rightleftharpoons cC + dD$ <p>The equilibrium constant <math>K_c</math> 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 expressed in terms of partial pressures. It is used for reactions involving gases.</p> <p>For a general reaction:</p> $aA + bB \rightleftharpoons cC + dD$ <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><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> and <math>K_b</math></b></p> <p><math>K_a</math> is the acid dissociation constant, and <math>K_b</math> is the base dissociation constant. They are used to measure the strength of acids and bases in aqueous solution.</p> <p>For an acid dissociation reaction:</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_a</math> and <math>K_b</math></b></p> <p>The relationship between <math>K_a</math> and <math>K_b</math> for a conjugate acid-base pair is given by the equation:</p> $K_a \times K_b = K_w$ <p>where <math>K_w</math> is the ion-product constant for water, which is <math>1.0 \times 10^{-14}</math> at 25°C.</p>	<p><b>Equilibrium Constant, <math>K_a</math> and <math>K_b</math></b></p> <p>The relationship between <math>K_a</math> and <math>K_b</math> for a conjugate acid-base pair is given by the equation:</p> $K_a \times K_b = K_w$ <p>where <math>K_w</math> is the ion-product constant for water, which is <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 border="1"> <thead> <tr> <th>Condition</th> <th>Effect on Equilibrium</th> <th>Example</th> </tr> </thead> <tbody> <tr> <td>Concentration</td> <td>Increasing the concentration of a reactant or product shifts the equilibrium in the direction that consumes the added substance.</td> <td>For the reaction <math>N_2 + 3H_2 \rightleftharpoons 2NH_3</math>, increasing the concentration of <math>N_2</math> shifts the equilibrium to the right, producing more <math>NH_3</math>.</td> </tr> <tr> <td>Pressure</td> <td>Increasing the pressure shifts the equilibrium in the direction that produces fewer moles of gas.</td> <td>For the reaction <math>N_2 + 3H_2 \rightleftharpoons 2NH_3</math>, increasing the pressure shifts the equilibrium to the right, producing more <math>NH_3</math>.</td> </tr> <tr> <td>Temperature</td> <td>Increasing the temperature shifts the equilibrium in the direction that absorbs heat (endothermic reaction).</td> <td>For the reaction <math>N_2 + 3H_2 \rightleftharpoons 2NH_3</math>, which is exothermic, increasing the temperature shifts the equilibrium to the left, producing less <math>NH_3</math>.</td> </tr> </tbody> </table>			Condition	Effect on Equilibrium	Example	Concentration	Increasing the concentration of a reactant or product shifts the equilibrium in the direction that consumes the added substance.	For the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$ , increasing the concentration of $N_2$ shifts the equilibrium to the right, producing more $NH_3$ .	Pressure	Increasing the pressure shifts the equilibrium in the direction that produces fewer moles of gas.	For the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$ , increasing the pressure shifts the equilibrium to the right, producing more $NH_3$ .	Temperature	Increasing the temperature shifts the equilibrium in the direction that absorbs heat (endothermic reaction).	For the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$ , which is exothermic, increasing the temperature shifts the equilibrium to the left, producing less $NH_3$ .
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**Minjie Lin**



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