



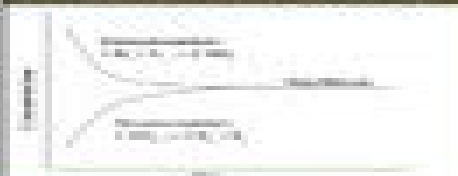
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CHEAT SHEET

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 $\rightleftharpoons$.

What is Equilibrium? 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. Equilibrium is a dynamic process, meaning that the reaction is still occurring, but the rates of the forward and reverse reactions are equal.	Equilibrium Constants K_c (Concentration): $K_c = \frac{[Products]}{[Reactants]}$ K_p (Pressure): $K_p = \frac{P_{Products}}{P_{Reactants}}$ K_a (Acid Dissociation): $K_a = \frac{[H^+][A^-]}{[HA]}$ K_b (Base Dissociation): $K_b = \frac{[OH^-][BH^+]}{[B]}$													
Equilibrium Constant, K_c For a general reaction: $aA + bB \rightleftharpoons cC + dD$ The equilibrium constant K_c is given by: $K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$ Where $[A]$, $[B]$, $[C]$, and $[D]$ are the molar concentrations of the reactants and products at equilibrium.	Equilibrium Constant, K_p <table><tr><th>Reaction</th><th>Equilibrium Constant</th></tr><tr><td>$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$</td><td>$K_p = \frac{P_{NH_3}^2}{P_{N_2} P_{H_2}^3}$</td></tr><tr><td>$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$</td><td>$K_p = \frac{P_{SO_3}^2}{P_{SO_2}^2 P_{O_2}}$</td></tr><tr><td>$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$</td><td>$K_p = \frac{P_{PCl_3} P_{Cl_2}}{P_{PCl_5}}$</td></tr></table>		Reaction	Equilibrium Constant	$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$	$K_p = \frac{P_{NH_3}^2}{P_{N_2} P_{H_2}^3}$	$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$	$K_p = \frac{P_{SO_3}^2}{P_{SO_2}^2 P_{O_2}}$	$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$	$K_p = \frac{P_{PCl_3} P_{Cl_2}}{P_{PCl_5}}$				
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Equilibrium Constant, K_a For a weak acid HA in water: $HA \rightleftharpoons H^+ + A^-$ The acid dissociation constant K_a is given by: $K_a = \frac{[H^+][A^-]}{[HA]}$	Equilibrium Constant, K_b <table><tr><th>Reaction</th><th>Equilibrium Constant</th></tr><tr><td>$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$</td><td>$K_b = \frac{[NH_4^+][OH^-]}{[NH_3][H_2O]}$</td></tr><tr><td>$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$</td><td>$K_b = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-][H_2O]}$</td></tr><tr><td>$CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$</td><td>$K_b = \frac{[HCO_3^-][OH^-]}{[CO_3^{2-}][H_2O]}$</td></tr></table>		Reaction	Equilibrium Constant	$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$	$K_b = \frac{[NH_4^+][OH^-]}{[NH_3][H_2O]}$	$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$	$K_b = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-][H_2O]}$	$CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$	$K_b = \frac{[HCO_3^-][OH^-]}{[CO_3^{2-}][H_2O]}$				
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Le Chatelier's Principle: The effect of changing conditions on equilibrium <table><tr><th>Condition</th><th>Effect on Equilibrium</th><th>Example</th></tr><tr><td>Concentration</td><td> - Increasing the concentration of a reactant shifts the equilibrium to the right. - Increasing the concentration of a product shifts the equilibrium to the left. </td><td>$Fe^{3+} + SCN^- \rightleftharpoons FeSCN^{2+}$ Adding more Fe^{3+} or SCN^- shifts the equilibrium to the right, forming more $FeSCN^{2+}$.</td></tr><tr><td>Pressure</td><td> - Increasing the pressure shifts the equilibrium to the side with fewer moles of gas. - Decreasing the pressure shifts the equilibrium to the side with more moles of gas. </td><td>$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ Increasing the pressure shifts the equilibrium to the right, forming more NH_3.</td></tr><tr><td>Temperature</td><td> - For an exothermic reaction, increasing the temperature shifts the equilibrium to the left. - For an endothermic reaction, increasing the temperature shifts the equilibrium to the right. </td><td>$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ This reaction is exothermic. Increasing the temperature shifts the equilibrium to the left, forming more SO_2 and O_2.</td></tr></table>			Condition	Effect on Equilibrium	Example	Concentration	- Increasing the concentration of a reactant shifts the equilibrium to the right. - Increasing the concentration of a product shifts the equilibrium to the left.	$Fe^{3+} + SCN^- \rightleftharpoons FeSCN^{2+}$ Adding more Fe^{3+} or SCN^- shifts the equilibrium to the right, forming more $FeSCN^{2+}$.	Pressure	- Increasing the pressure shifts the equilibrium to the side with fewer moles of gas. - Decreasing the pressure shifts the equilibrium to the side with more moles of gas.	$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ Increasing the pressure shifts the equilibrium to the right, forming more NH_3 .	Temperature	- For an exothermic reaction, increasing the temperature shifts the equilibrium to the left. - For an endothermic reaction, increasing the temperature shifts the equilibrium to the right.	$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ This reaction is exothermic. Increasing the temperature shifts the equilibrium to the left, forming more SO_2 and O_2 .
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