



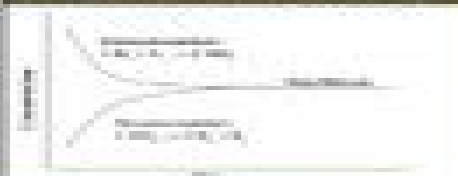
CHEMISTRY

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 $\rightleftharpoons$.

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>$2NO_2(g) \rightleftharpoons N_2O_4(g)$ Adding more NO_2 shifts the equilibrium to the right, producing more N_2O_4.</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, producing 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, producing 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.	$2NO_2(g) \rightleftharpoons N_2O_4(g)$ Adding more NO_2 shifts the equilibrium to the right, producing more N_2O_4 .	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, producing 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, producing more SO_2 and O_2 .
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