During Class Invention #

Name(s) with Lab section in Group

Strong Acids and the Equilibrium Constant

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1. The chemical equation which describes how HCl dissociates in aqueous solution is,

$$HCl(aq) \rightleftharpoons H^+(aq) + Cl^-(aq)$$

a. In the demonstration performed earlier, the initial concentration of HCl is 0.100 M. In the space provided below (**ICE** Table), enter the initial concentration of HCl, H^+ and Cl^-. Based on the measured pH of this solution, calculate and enter the equilibrium concentration of H^+.

	$HCl(aq)$	$\rightleftharpoons$ $H^+(aq)$	+ $Cl^-(aq)$
Initial Concentrations	**0.100 M**	**~ 0 M**	**0 M**
Change	**-0.1 M**	**+0.1 M**	**+0.1 M**
Equilibrium Concentrations	**0 M**	**+0.1 M**	**+0.1 M**

b. Calculate the change in $[H^+]$.

The $[H^+]$, based on the measured pH, is 0.1 M. Actually, the initial $[H^+]_0$ is 1 x 10^{-7} M. But, for our purposes, we can assume the initial $[H^+]$ is 0 M. So the change in $[H^+]$ is 0.100 M.

c. Using the balanced chemical equation and the calculated change in $[H^+]$, calculate the change in HCl and Cl^-.

Since the stoichiometry in the balanced chemical equation is 1 : 1, the change in the [HCl] must be equal to change in $[H^+]$. Therefore,

$$\Delta[H^+] = -\Delta[HCl] \text{ and } \Delta[H^+] = \Delta[Cl^-] = 0.100 \text{ M}$$

d. Calculate the equilibrium concentration of HCl and Cl^-.

$$[HCl]_{eq} = [HCl]_0 - \Delta[HCl] = 0.100 \text{ M} - 0.100 \text{ M} = \sim 0 \text{ M}$$

$$[Cl^-]_{eq} = [Cl^-]_0 + \Delta[Cl^-] = 0 \text{ M} + 0.100 \text{ M} = 0.100 \text{ M}$$

e. Estimate the equilibrium constant for the dissociation of $HCl(aq)$.

$$K_a = \frac{[H^+][Cl^-]}{[HCl]} = \frac{0.1 \text{ M} \cdot 0.1 \text{ M}}{\sim 0 \text{ M}} = \text{very large number}$$

The value is approximately 1 x 10^{12}.