Chem 1515
During Class Invention
Spring 2007
Arrhenius Acids [H^+], [OH^-], pH and pOH

Name___________________________
TA Name ______________________
Lab Section #______

1. a. Hydrochloric acid, HCl(aq), breaks up into ions (a cation and an anion) in a water solution. Write a chemical equation that represents this process.

$$\mathbf{HCl}(aq) \rightarrow \mathbf{H^+}(aq) + \mathbf{Cl^-}(aq)$$

   b. Nitric acid, HNO_3(aq), breaks up into ions (a cation and an anion) in a water solution. Write a chemical equation that represents this process.

$$\mathbf{HNO_3}(aq) \rightarrow \mathbf{H^+}(aq) + \mathbf{NO_3^-}(aq)$$

   c. Write a definition describing what happens to acids when they interact with water.

An acid is a substance that when dissolved in water ionizes to form hydrogen ion, H^+*(aq)*.

2. a. Sodium hydroxide, NaOH(aq), breaks up into ions (a cation and an anion) in a water solution. Write a chemical equation that represents this process.

$$\mathbf{NaOH}(aq) \rightarrow \mathbf{Na^+}(aq) + \mathbf{OH^-}(aq)$$

   b. Potassium hydroxide, KOH(aq), breaks up into ions (a cation and an anion) in a water solution. Write a chemical equation that represents this process.

$$\mathbf{KOH}(aq) \rightarrow \mathbf{K^+}(aq) + \mathbf{OH^-}(aq)$$

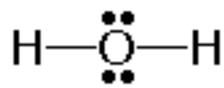
   c. NaOH and KOH are classified as bases. Write a definition describing what happens to bases when they interact with water.

A base is a substance that when dissolved in water ionizes to form hydroxide ion, OH^-*(aq)*.

3. Defining acids and bases in terms of the characteristic ions that are released in water solution is the Arrhenius Theory of acids and bases. In the space below, list some examples of Arrhenius acids and Arrhenius bases.

Arrhenius Acids	Arrhenius Bases
HCl	**$NaOH$**
H_2SO_4	**KOH**
$HC_2H_3O_2$	**$Ba(OH)_2$**
HBr	**$Ca(OH)_2$**
$HClO_4$	**NH_3**

4. a. Draw the Lewis structure for water, H_2O.



b. Describe how water might break up into ions. Write a chemical equation for this process. How is water related to Arrhenius acids and bases?

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

Or

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

Water breaks up into the ions that are characteristic of acids, H^+*(aq)* and bases, OH^-*(aq)*.

c. Unlike HCl only a small fraction of water molecules break up into ions. The equilibrium constant for water dissociating into ions is 1.0 x 10^{-14}. Write an equilibrium expression for the dissociation of water. What would be the concentration of each of the ions in a pure sample of water?

	$H_2O(l)$ $\rightleftharpoons$	$H^+(aq)$ +	$OH^-(aq)$	
Initial	----	0	0	
Change	----	+x	+x	$x = [H_2O]_{reacted}$
Equilibrium	----	0+x	0+x	

$$K_w = [H^+][OH^-]$$
$$1.0 \times 10^{-14} = (x)(x)$$
$$1.0 \times 10^{-7}\ M = x$$

5. a. pH is a measure of the $[H^+]$. The pH of pure water =7. Write an equation that relates pH to $[H^+]$ by comparing 7 with the $[H^+]$ of pure water.

$$\mathbf{pH = -\log [H^+]}$$

   b. Use LeChatelier's Principle to predict what would happen to the $[H^+]$ and $[OH^-]$ of a water sample if you added HCl to it. What would happen to the value of the pH.

   Adding HCl to a sample of water will result in the $[H^+]$ increasing. If the $[H^+]$ increases than the $[OH^-]$ must decrease so the product of $[H^+]\cdot[OH^-] = 1.0 \times 10^{-14}$

   c. Use LeChatelier's Principle to predict what would happen to the $[H^+]$ and $[OH^-]$ of a water sample if you added NaOH to it. What would happen to the value of the pH.

   Adding NaOH to a sample of water will result in the $[H^+]$ decreasing. If the $[H^+]$ decreases than the $[OH^-]$ must increase so the product of $[H^+]\cdot[OH^-] = 1.0 \times 10^{-14}$

   d. pOH is a measure of the $[OH^-]$. The pOH of pure water =7. Write an equation that relates pOH to $[OH^-]$ by comparing 7 with the $[OH^-]$ of pure water.

$$\mathbf{pOH = -\log [OH^-]}$$

6. a. The $[H^+]$ in a particular aqueous solution is 1.0×10^{-4} M. Calculate the $[OH^-]$ for this solution.

$$K_w = [H^+][OH^-]$$
$$1.0 \times 10^{-14} = (1.0 \times 10^{-4}\ \text{M})[OH^-]$$
$$\frac{1.0 \times 10^{-14}}{(1.0 \times 10^{-4}\ \text{M})} = [OH^-]$$
$$1.0 \times 10^{-10}\ \text{M} = [OH^-]$$

   b. The $[OH^-]$ in a particular aqueous solution is 1.0×10^{-5} M. Calculate the $[H^+]$ for this solution.

$$K_w = [H^+][OH^-]$$
$$1.0 \times 10^{-14} = [H^+](1.0 \times 10^{-5}\ \text{M})$$
$$\frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}\ \text{M}} = [H^+]$$
$$1.0 \times 10^{-9}\ \text{M} = [H^+]$$

c. The $[H^+]$ in a particular aqueous solution is 6.0 M. Calculate the $[OH^-]$ for this solution.

$$K_w = [H^+][OH^-]$$
$$1.0 \times 10^{-14} = (6.0\ M)[OH^-]$$
$$\frac{1.0 \times 10^{-14}}{6.0\ M} = [OH^-]$$

1.7×10^{-15} M = $[OH^-]$ note it is possible for the $[H^+]$ or the $[OH^-]$ to be smaller than 1×10^{-14} M.

7. a. Calculate the pH and pOH of a solution with a $[H^+]$ 3.68×10^{-8} M.

$$pH = -\log[H^+]$$
$$pH = -\log[3.68 \times 10^{-8}]$$
$$pH = -(-7.43)$$
$$pH = 7.43$$

$$pH + pOH = 14$$
$$7.43 + pOH = 14$$
$$pOH = 14 - 7.43$$
$$pOH = 6.57$$

b. Calculate the the $[H^+]$ and $[OH^-]$ of a solution with a pH = 4.22.

$$pH = -\log[H^+]$$
$$4.22 = -\log[H^+]$$
$$-4.22 = \log[H^+]$$
$$10^{-4.22} = 10^{\log[H+]}$$
$$6.03 \times 10^{-5}\ M = [H^+]$$

$$K_w = [H^+][OH^-]$$
$$1.0 \times 10^{-14} = (6.03 \times 10^{-5})[OH^-]$$
$$\frac{1.0 \times 10^{-14}}{6.03 \times 10^{-5}} = [OH^-]$$
$$1.66 \times 10^{-10}\ M = [OH^-]$$