During Class Invention #

Name(s) with Lab section in Group

Arrhenius Acids, $[H^+]$, $[OH^-]$ pH and pOH

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1a. Define the terms *Arrhenius acid* and *Arrhenius base*.

An acid is a substance which, when dissolved in water, increases the concentration of hydrogen ion, $H^+(aq)$. For example,

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

A base is a substance which, when added to water, increases the concentration of hydroxide ion, $OH^-(aq)$. For example,

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

b. Write a chemical equation that describes the behavior of an Arrhenius acid in water.

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

c. Write a chemical equation that describes the behavior of an Arrhenius base in water.

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

2. In the space below, list some examples of Arrhenius acids and Arrhenius bases.

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

2a. Write the autoionization reaction for water and the equilibrium expression for the autoionization reaction.

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

$$\mathbf{K_w = [H^+][OH^-]}$$

or another way

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

$$\mathbf{K_w = [H_3O^+][OH^-]}$$

b. What is the magnitude of the equilibrium constant at 25 ˚C for the autoionization reaction of water?

$$K_w = 1.0 \times 10^{-14}$$

c. What are the concentrations of H^+ and OH^- at 25 ˚C in pure water?

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

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

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

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

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

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

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

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

3a. Define *pH* and *pOH* for aqueous solutions of acids or bases.

$pH = -\log[H^+]$
$pOH = -\log[OH^-]$

$pH + pOH = 14$
For neutral aqueous solutions, the pH = 7.

b. Indicate the range of pH which characterizes an acidic solution and the range which characterizes a basic solution.

pH Range	
0 – 6.99	**acidic**
7	**neutral**
7.01 - 14	**basic**

c. Calculate the pH and pOH of a solution with a $[H^+] = 3.68 \times 10^{-8}$ M.

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

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

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

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

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