

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*.

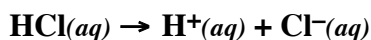
**An acid is a substance which, when dissolved in water, increases the concentration of hydrogen ion,  $H^+(aq)$ . For example,**



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



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



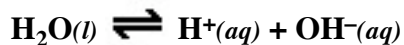
c. Write a chemical equation that describes the behavior of an Arrhenius base in water.



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

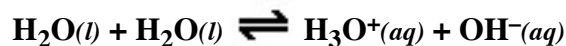
Arrhenius Acids	Arrhenius Bases
HCl	NaOH
H <sub>2</sub> SO <sub>4</sub>	KOH
HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>	Ba(OH) <sub>2</sub>
HBr	Ca(OH) <sub>2</sub>
HClO <sub>4</sub>	

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



$$K_w = [H^+][OH^-]$$

or another way

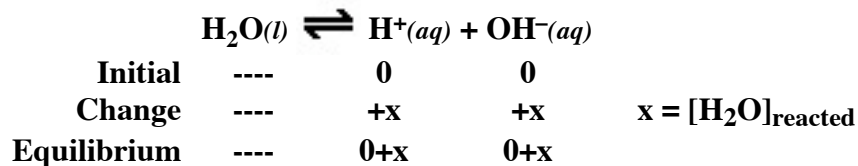


$$K_w = [H_3O^+][OH^-]$$

- 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<sup>+</sup> and OH<sup>-</sup> at 25 °C in pure water?



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

$$1.0 \times 10^{-14} = (x)(x)$$

$$1.0 \times 10^{-7} \text{ M} = x$$

- d. The [H<sup>+</sup>] in a particular aqueous solution is 1.0 x 10<sup>-4</sup> M. Calculate the [OH<sup>-</sup>] for this solution.

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

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

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

$$1.0 \times 10^{-10} \text{ M} = [\text{OH}^-]$$

- e. The [OH<sup>-</sup>] in a particular aqueous solution is 1.0 x 10<sup>-5</sup> M. Calculate the [H<sup>+</sup>] of this solution.

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

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

$$\frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5} \text{ M}} = [\text{H}^+]$$

$$1.0 \times 10^{-9} \text{ M} = [\text{H}^+]$$

- f. The [H<sup>+</sup>] in a particular aqueous solution is 6.0 M. Calculate the [OH<sup>-</sup>] of this solution.

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

$$1.0 \times 10^{-14} = (6.0 \text{ M})[\text{OH}^-]$$

$$\frac{1.0 \times 10^{-14}}{6.0 \text{ M}} = [\text{OH}^-]$$

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

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.

<b>pH Range</b>	
<b>0 – 6.99</b>	<b>acidic</b>
<b>7</b>	<b>neutral</b>
<b>7.01 - 14</b>	<b>basic</b>

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  $pH = 4.22$ .

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

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