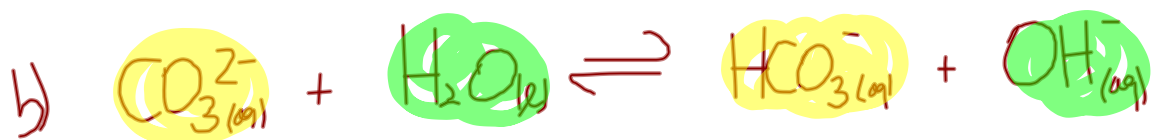


## Homework - #3-5,7,8

Arrhenius: Acids  $\rightarrow$   $H^+$   
Bases  $\rightarrow$   $OH^-$

Bronsted-Lowry: Acids  $\rightarrow$   $H^+$  proton donor  
Bases  $\rightarrow$  proton acceptors



# Water Equilibrium

Conductivity is due to the presence of ions. For water:



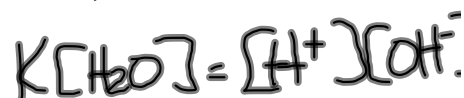
- therefore  $K = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$  is very small

- slight conductivity shows that equilibrium greatly favors water molecules (less than 2  $\text{H}^+$  per billion water)

- therefore the concentration of water in pure water and in dilute aqueous solutions is essentially constant and can be combined with the equilibrium constant to produce a new constant called the *Ion Product Constant*

## Ionization Constant for water (ion product constant)

$$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ at SATP}$$



Since  $[\text{H}^+]$  and  $[\text{OH}^-]$  are found in 1:1 ratio  
 $(\text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}^+_{(aq)} + \text{OH}^-_{(aq)})$

$[\text{H}^+_{(aq)}] = [\text{OH}^-_{(aq)}] = 1.0 \times 10^{-7} \text{ mol/L}$  in **neutral** solutions.

Arrhenius's Theory - acid is a substance that ionizes water to produce  $\text{H}^+$  ions.

- additional ions produced by the acid increases the  $\text{H}^+$  concentration in the water. (more acid, more  $\text{H}^+$ )

**Therefore acids always have a  $[\text{H}^+] > 10^{-7} \text{ mol/L}$**

**Basic solutions produce a  $[\text{OH}^-]$  greater than  $10^{-7} \text{ mol/L}$**

$K_w$  can be used to calculate either  $[\text{H}^+]$  or  $[\text{OH}^-]$

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

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

# pH and pOH

$$\text{pH} = -\log[\text{H}^+_{(\text{aq})}]$$

$$\text{pOH} = -\log[\text{OH}^-_{(\text{aq})}]$$

$$[\text{H}^+_{(\text{aq})}] = 10^{-\text{pH}}$$

$$[\text{OH}^-_{(\text{aq})}] = 10^{-\text{pOH}}$$

Ex. Calculate the pH of a solution where  $[\text{H}^+_{(\text{aq})}] = 3.24 \times 10^{-4} \text{M}$ .

$$\begin{aligned}\text{pH} &= -\log[\text{H}^+_{(\text{aq})}] \\ \text{pH} &= -\log[3.24 \times 10^{-4}] \\ \text{pH} &= 3.489\end{aligned}$$

Ex. Calculate the concentration of hydroxide ions in a solution with a pOH of 10.14.

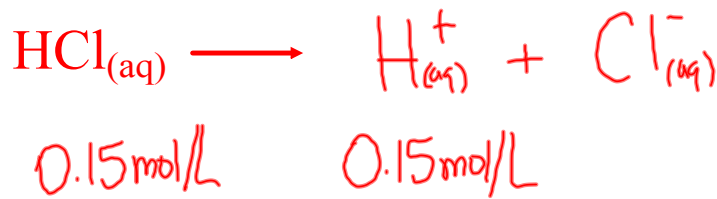
$$\begin{aligned}[\text{OH}^-_{(\text{aq})}] &= 10^{-\text{pOH}} \\ [\text{OH}^-_{(\text{aq})}] &= 10^{-10.14} \\ [\text{OH}^-_{(\text{aq})}] &= 7.2 \times 10^{-11} \text{M}\end{aligned}$$

$10^x$  ( $= 10.14$ )

# Strong Acids

Calculate the concentration of the hydroxide ions, pH and pOH of a 0.15 mol/L solution of hydrochloric acid at 25°C.

**\*Strong acids will always completely ionize\***



pH = ? ✓  
pOH = ?  
[OH<sup>-</sup>] = ?

$$\text{pH} = -\log[\text{H}_{(aq)}^{+}]$$
$$\text{pH} = -\log[0.15]$$

pH = 0.82

$$K_w = [\text{H}_{(aq)}^{+}][\text{OH}_{(aq)}^{-}]$$

$$[\text{OH}_{(aq)}^{-}] = \frac{K_w}{[\text{H}_{(aq)}^{+}]}$$

$$[\text{OH}_{(aq)}^{-}] = \frac{1.0 \times 10^{-14}}{0.15 \text{ mol/L}}$$

[OH<sup>-</sup>]<sub>(aq)</sub> = 6.7 × 10<sup>-14</sup> M

$$\text{pOH} = -\log[\text{OH}_{(aq)}^{-}]$$

$$\text{pOH} = -\log[6.7 \times 10^{-14}]$$

pOH = 13.18

**#11,12 p.599**

**#13,14 p.600**

**#15,16 p.601**

**#17-21 p.604**