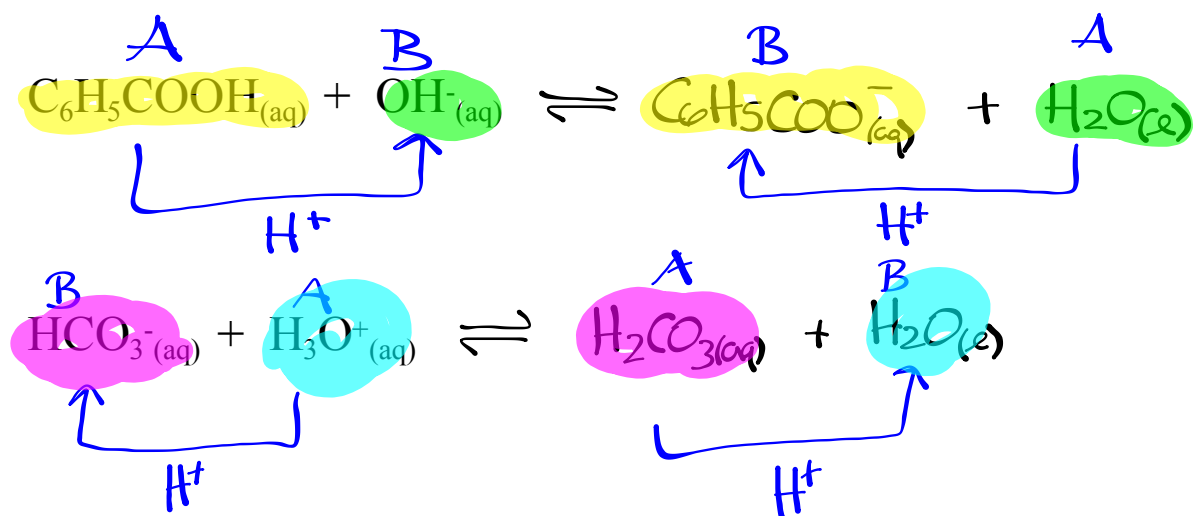


# Warm Up

Predict the products for the following reaction, and identify each reactant as an acid or a base.



## Revised Arrhenius Theory

Acids  $\text{H}^+$   
Bases  $\text{OH}^-$

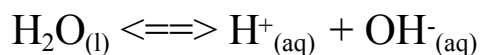
(acid/base  $\neq$  water)  
bases  $\rightarrow \text{Na}_2\text{CO}_3, \text{NH}_3$

## Bronsted-Lowry Theory

Acids  $\rightarrow \text{H}^+$  donors  
Bases  $\rightarrow \text{H}^+$  acceptors

## Water Equilibrium

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

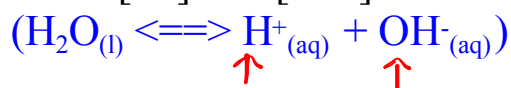


- therefore  $K = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$  is very small
- slight conductivity shows that equilibrium greatly favors water molecules (less than 2  $\text{H}^+$  per billion water)
- therefore the concentration of water in pure water and in dilute aqueous solutions is essentially 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)

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

Since  $[\text{H}^+]$  and  $[\text{OH}^-]$  are found in 1:1 ratio



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

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  $[\text{OH}^-]$

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

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

$$K = \frac{[H^+][OH^-]}{[H_2O]}$$

$$\underbrace{K[H_2O]} = [H^+][OH^-]$$

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

# pH and pOH

acidic

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

basic

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

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

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

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

$$\text{pH} + \text{pOH} = 14.00$$

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

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

$$\text{pH} = -\log[3.24 \times 10^{-4}]$$

$$\boxed{\text{pH} = 3.489}$$

$$3.24 \left[ \begin{array}{c} \boxed{\text{Exp}} \\ \boxed{\text{EE}} \end{array} \right] = \boxed{4}$$

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

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

$$[\text{OH}^-_{(\text{aq})}] = 10^{-10.14}$$

$$\boxed{[\text{OH}^-_{(\text{aq})}] = 7.2 \times 10^{-11} \text{M}}$$

$$\boxed{10^x}$$

$$10^{(-10.14)}$$

