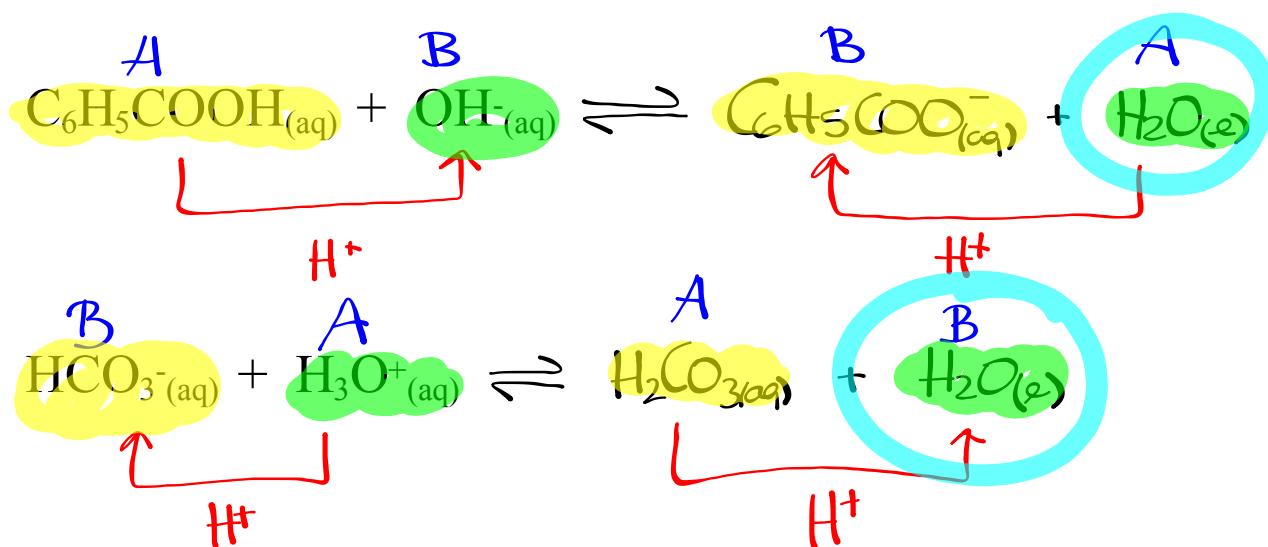


Warm Up

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



Conjugate acid-base pairs

two substances that differ by a proton

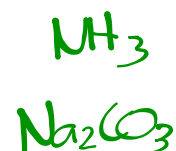
Ex. H_2O and H_3O^+

NH_4^+ and NH_3

Revised Arrhenius Theory

ACID - H^+ ions

BASE - OH^- ions



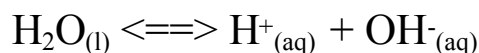
Bronsted-Lowry Theory

ACID - H^+ (proton) donors

BASE H^+ (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 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)

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

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

[H⁺_(aq)] = [OH⁻_(aq)] = 1.0 × 10⁻⁷ mol/L in **neutral** solutions.

Arrhenius's Theory - acid is a substance that ionizes water to produce H⁺ ions.

- additional ions produced by the acid increases the H⁺ concentration in the water. (more acid, more H⁺)

Therefore acids always have a [H⁺] > 10⁻⁷ mol/L

Basic solutions produce a [OH⁻] greater than 10⁻⁷ mol/L

K_w can be used to calculate either [H⁺] or [OH⁻]

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

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



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

$$K[\text{H}_2\text{O}] = [\text{H}^+][\text{OH}^-]$$

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

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

$$[\text{H}^+]^2 = 1.0 \times 10^{-14}$$

$$[\text{H}^+] = 1.0 \times 10^{-7}$$

pH and pOH

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

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

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

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

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

Ex. Calculate the pH of a solution where

$$[\text{H}^+_{(\text{aq})}] = 3.24 \times 10^{-4} \text{ M}$$

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

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

$$\text{pH} = 3.489$$

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

$$[\text{OH}^-_{(\text{aq})}] = \frac{1.0 \times 10^{-14}}{3.24 \times 10^{-4}}$$

$$[\text{OH}^-_{(\text{aq})}] = 3.09 \times 10^{-11} \text{ M}$$

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

$$\text{pOH} = -\log[3.09 \times 10^{-11}]$$

$$\text{pOH} = 10.510$$

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



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

$$\text{pH} = -\log[0.15]$$

$$\boxed{\text{pH} = 0.82}$$

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

$$\text{pOH} = 14.00 - 0.82$$

$$\boxed{\text{pOH} = 13.18}$$

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

$$[\text{OH}^-_{(aq)}] = 10^{-13.18}$$

$$\boxed{[\text{OH}^-_{(aq)}] = 6.6 \times 10^{-14} \text{ M}}$$