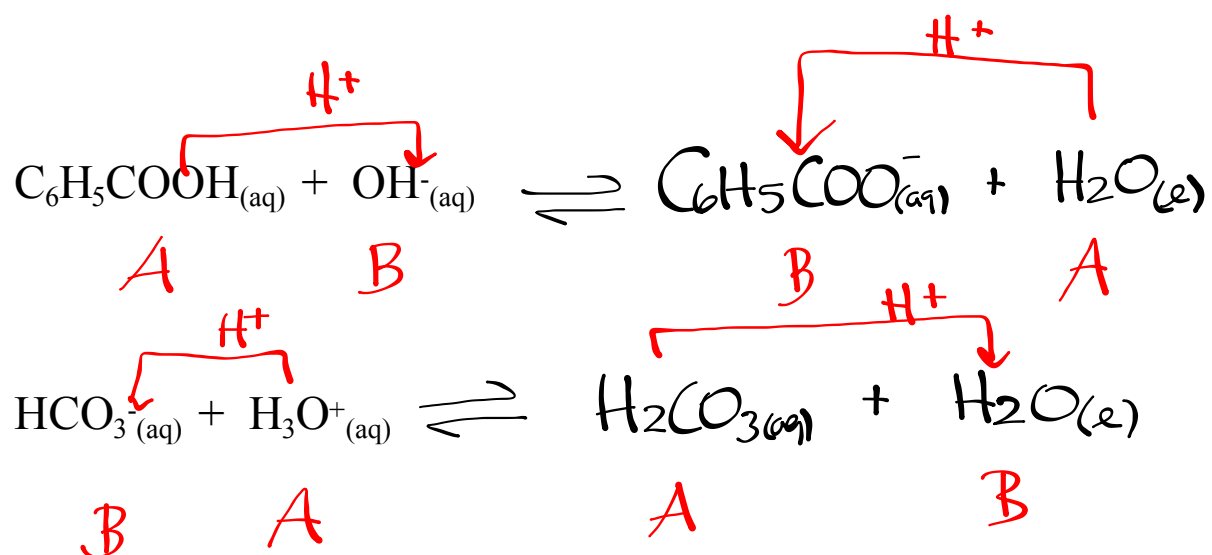


Warm Up

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



^{Revised} Arrhenius Theory

Acid - H^+

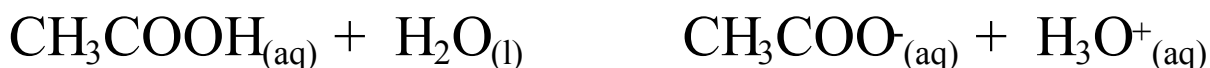
Base - OH^-

Bronsted-Lowry Theory

Acid - H^+ donors

Base - H^+ acceptors

Conjugate Acid-Base Pairs



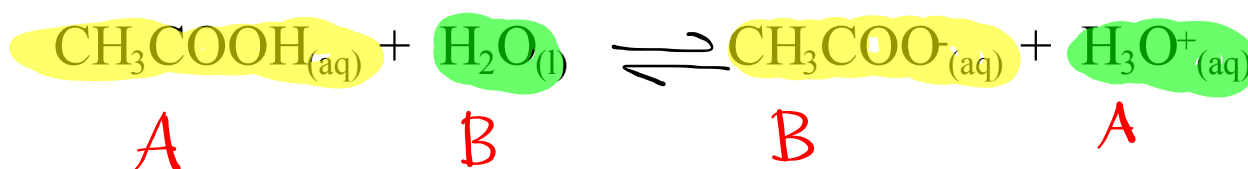
Acid-Base reactions are at equilibrium !

(Look at forward reaction and reverse reaction)

- Every acid-base reaction at equilibrium has two acids and two bases.
- Acid on 'product' side is formed by addition of proton to base on 'reactant' side
- Base on 'product' side is formed by removal of a proton from acid on 'reactant' side

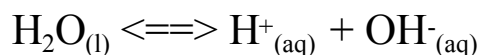
Conjugate acid-base pair

A pair of substances that differ by only a proton



Water Equilibrium

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

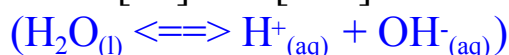


- therefore $K = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$ is very small
- slight conductivity shows that equilibrium greatly favors water molecules (less than 2 H⁺ per billion water)
- therefore the concentration of water in pure water and in dilute aqueous solutions is essentially constant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



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

Arrhenius's Theory - acid is a substance that ionizes water to produce 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⁻]

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

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

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

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

Hand-drawn calculator keypad showing the calculation of pH: EE, 3.24, EXP, =, 4.

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

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

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

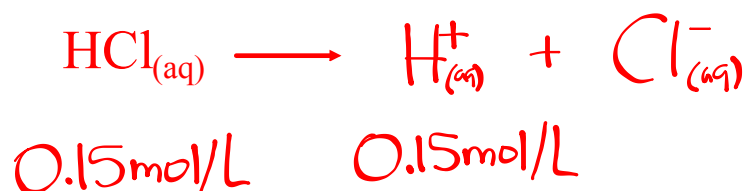
$$[\text{OH}^-_{(\text{aq})}] = 7.2 \times 10^{-11} \text{M}$$

Hand-drawn calculator keypad showing the calculation of hydroxide ion concentration: 10, y*, (-), 10.14, ^.

Strong Acids

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

Strong acids will always completely ionize



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

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

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

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

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

$$\boxed{\text{pOH} = 13.17}$$

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

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

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

$$\boxed{[\text{OH}^-_{(\text{aq})}] = 6.7 \times 10^{-14} \text{ M}}$$

Strong Bases (Ionic Hydroxides)

Calculate the hydrogen ion concentration in a 0.25 mol/L solution of barium hydroxide.

