

IONIC EQUILIBRIA OF ACIDS and BASES (Ch.14 (sec.1-5))

THEORY	SOLVENT RESTRICTIONS	ACID	BASE
Arrhenis	Aqueous only	H ₃ O ⁺ (H ⁺)	OH ⁻
Bronsted-Lowry	Protonic solvents	Proton donor	Proton acceptor
Lewis	None	e-pair acceptor	e- pair donor

B/L Theory is followed mostly - where conjugate base (CB) - accepts a p⁺ and conjugate acid - donates a p⁺

Water undergoes self-ionization:



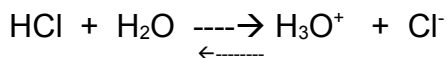
$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

So each concentration (M) of hydronium and hydroxide = 1.0×10^{-7} M

Remember [H⁺] = [H₃O⁺]

Also strong electrolytes ionize totally in solution and weak electrolytes, only partially.

Ex: HCl (monoprotic acid) will ionize totally to:



****because the moles of water consumed in the formation of hydronium are negligible compared to the # of moles present, M is constant, so no need for it in the K_A.

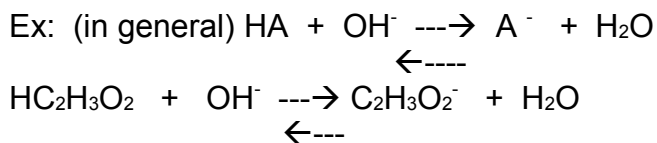
K_A is the acid dissociation constant. It equals the concentration of the products/ concentration of reactants.

The higher the K_A, then the stronger the acid and the more it dissociates. Same thinking for bases. (K_B)

- There are, however, many strong acids that don't have an equilibrium constant. Why? Concentrations can be found from the molar concentration of the solution.

So, most of the time when equilibrium constants are found for acids or bases, it is because they are weak!

Weak acids will only slightly ionize in water:
(Ch.14.2 - p.661)



REVIEW OF pH (Ch14 p.666)

Neutral	[H ⁺] = [OH ⁻]	pH = -log [H ⁺]
Acidic	[H ⁺] > [OH ⁻]	pOH = -log[OH ⁻]
Basic	[H ⁺] < [OH ⁻]	

$$K_w = [H_3O^+] + [OH^-] = 1.0 \times 10^{-14}$$

(refer to self-ionization equation)

Calculating the pH of Strong Acid Solutions

In attacking acid-base equilibrium problems, writing the major species in the solution is VITAL!

Example 14-7 p.670

1. Calculate the pH of .10 M HNO₃.

It is a strong acid and the major species are: H⁺, NO₃⁻, & H₂O

The M of the acid is pretty much 0, since the acid completely dissociates. The hydroxide ion concentration will be very small too because the hydronium ions from the acid drives equilibrium to the left. So the sources of hydronium ion are from HNO₃ (.10 M) and from H₂O.

The H⁺ ions contributed by the self-ionization of water is so small, it can be ignored compared to the ions coming from the .10 M HNO₃.

So, [H⁺] = .10 M and pH = -log[.10] = 1.00 (very acidic)

2. Calculate the pH of 1.0 x 10⁻¹⁰ M HCl.

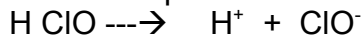
The amount of HCl in solution is very small, so the only major species is H₂O. pH = 7.00

Calculating the pH of Weak Acids (starts with Ch 14.5)

Calculate the [H⁺] and % ionization of .0092 M H ClO is given:

(like 14.10- p.678- good supplemental article on p. 679)

Write the species involved:



←---

$$K_A = \frac{[H^+][ClO^-]}{[HClO]} = 3.5 \times 10^{-8} \quad (\text{found on a reference table})$$

The amount of H⁺ and OH⁻ in solution that is from H₂O is too small and can be ignored.

	[HClO]	[H ⁺]	[ClO ⁻]
Initial	.0092	0	0
Equilibrium	.0092-x	X	X

$$X = [H^+] = [ClO^-]$$

$$[HClO] = .0092 - x$$

$$K_A = 3.5 \times 10^{-8} = \frac{x^2}{.0092 - x}$$

*** if the concentration of ionizing substance and K_A differ by 5% or more, than x can be neglected in the denominator. (otherwise, you have to plug it in to your favorite of all time...the quadratic formula!!!)

$$So, 3.5 \times 10^{-8} = \frac{x^2}{.0092}$$

$$x = 1.8 \times 10^{-5} M = [H^+] = [ClO^-]$$

$$\frac{1.8 \times 10^{-5} M}{.0092 M} \times 100 = \% \quad \text{part/total} = \% \text{ ionized in a WA}$$