Chapter 15 Weak Acids & Bases

Equilibrium Expression for Acids

In general terms, the reaction is:

HA	+	H_2O	<=>	H_3O^+	+	A^-
Weak Bronsted acid						Conjugate base

$$K_c = \frac{[H_3O^+][A^-]}{[H_2O][HA]}$$

$$K_c[H_2O] = \frac{[H_3O^+][A^-]}{[HA]}$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

Sometimes reaction is shortened to: HA <=> H^+ + A^-

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Better way to classify acid strength is by K_a.

$K_a < 10^{-3}$	Weak acid
$K_a = 1$ to 10^{-3}	Moderate acid
$K_a > 1$	Strong acid

For the general reaction of

HA	+	H_2O	<=>	H_3O^+	+	A^-
Bronsted acid						Conjugate base

the "A" denotes any acid; that is, particle with ability to donate a proton, so "A" could be:

Examples:

neutral molecule	$HC_2H_3O_2 + H_2O \Leftrightarrow H_3O^+ + C_2H_3O_2^-$
positive ion	$NH_4^+ + H_2O \Leftrightarrow H_3O^+ + NH_3$
negative ion	$HSO_4^- + H_2O \Leftrightarrow H_3O^+ + SO_4^{2-}$

The stronger the acid, the weaker is the conjugate base and vice-versa.

Calculations - Weak Acid Equilibria

Example 1:
Lactic acid ($C_3H_6O_3$) is a weak monoprotic acid. Determine the K_a for lactic acid if a 0.10M solution of the acid has a pH of 2.43.

Steps to filling in ICE chart:

(A) Concentration of acid is the initial concentration (HA).

(B) Pure water has an initial concentration of 1.0 x10^{-7}M H_3O^+ and 1.0 x10^{-7}M OH^- but generally this value is so small relative to amount produced by acid or base that considered negligible. However, NOT negligible in very dilute solutions.

(C) pH values are equilibrium concentrations.

	$C_3H_6O_3$ + H_2O ⇔	$C_3H_5O_3^-$ +	H_3O^+
I	0.10 M	0	~ 0
C	– x	+ x	+ x
E	0.10 – x	x	x = 10^{-pH}

pH = 2.43
$[H_3O^+] = 10^{-pH} = 10^{-2.43} = 0.0037$ M
so x = 0.0037 M

$$K_a = \frac{[C_3H_5O_3^-][H_3O^+]}{[C_3H_6O_3]}$$

$$K_a = \frac{x^2}{(0.10\text{-}x)} = \frac{(0.0037)^2}{(0.10\text{-}0.0037)} = 1.4x10^{-4}$$

Example 2:
Calculate the pH, % ionization, and the concentrations of all species (HF, F^-, H_3O^+, OH^-) present in 0.050M HF solution. $K_a = 3.5 \text{ x } 10^{-4}$

	HF + H_2O ⇔	F^- +	H_3O^+
I	0.050 M	0	~ 0
C	– x	+ x	+ x
E	0.050 – x	x	x

$$K_a = \frac{[F^-][H_3O^+]}{[HF]}$$

$$3.5 \text{ x } 10^{-4} = \frac{x^2}{(0.050\text{-}x)}$$

$$0 = \frac{x^2}{(0.050\text{-}x)} - 3.5 \text{ x } 10^{-4}$$ Window x= 0 to 0.050

Intersection
$x = 0.0040$ M

$[H_3O^+] = [F^-] = x = 0.0040$ M

$[HF] = 0.050 - x = 0.050\ M - 0.0040\ M = 0.046$ M

$pH = -\log([H_3O^+]) = -\log(x) = -\log(0.0040) = 2.40$

$pOH = 14.00 - pH = 14.00 - 2.40 = 11.60$

$$\%\text{ ionization} = \frac{[H_3O^+]}{[HF]_{INITIAL}} = \frac{x}{[HF]} = \frac{0.0040\ M}{0.050\ M} \times 100 = 8.0\%$$

Calculations - Weak Base Equilibria

Calculate the pH of a 0.015M NH_3 solution. $K_b = 1.8 \times 10^{-5}$
$NH_3\ (aq) + H_2O\ (l) \Leftrightarrow NH_4^+\ (aq) + OH^-\ (aq)$

	NH_3	+	H_2O	⇔	NH_4^+	+	OH^-
I	0.015 M				0		~ 0
C	– x				+ x		+ x
E	0.015 – x				x		x

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$1.8 \times 10^{-5} = \frac{x^2}{(0.015-x)}$$

$$0 = \frac{x^2}{(0.015-x)} - 1.8 \times 10^{-5}$$ Window x= 0 to 0.015

Intersection
$x = 5.1 \times 10^{-4}$ M

$[OH^-] = x = 5.1 \times 10^{-4}$ M
$pOH = -\log([OH^-]) = -\log(x) = -\log(5.1 \times 10^{-4}) = 3.29$

$pH = 14.00 - pOH = 14.00 - 3.29 = 10.71$

Calculations - Weak Acid/Base Equilibria

pH of a salt solution

Calculate the pH of a 0.15 M KCN solution.

K^+	CN^-
~~A~~	B

CN^- is the conjugate base of the weak acid HCN. $K_a = 6.2 \times 10^{-10}$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-10}} = 1.6 \times 10^{-5}$$

	CN^- + H_2O ⇔	HCN +	OH^-
I	0.15 M	0	~ 0
C	– x	+ x	+ x
E	0.15 – x	x	x

$$K_b = \frac{[HCN][OH^-]}{[CN^-]}$$

$$1.6 \times 10^{-5} = \frac{x^2}{(0.15 - x)}$$

$$0 = \frac{x^2}{(0.15 - x)} - 1.6 \times 10^{-5}$$ Window x = 0 to 0.15

Intersection

x = 0.0015 M

$[OH^-] = x = 0.0015$ M

pOH = –log ($[OH^-]$) = –log(x) = –log(0.0015) = 2.81

pH = 14.00 – pOH = 14.00 – 2.81 = 11.19