Solve K_a and K_b Problems Using ICE Methods

There are 3 basic types of problems.

A. Calculate K_a, K_b values from weak acids/weak bases.
B. Calculate values of $[H^+]$ and pH of a solution of a weak acid/base.
C. Calculate pH of salt solutions.

A. Calculate K_a, K_b values from weak acids/weak bases.

Lactic acid is a weak acid. The equilibrium equation is:

$$HC_3H_5O_3 \rightarrow H^+ + + C_3H_5O_3^-$$

In a .100m solution, the pH = 2.44. Find H^+ concentration. Find K_a using H^+ concentration.

Perform ICE:

	$HC_3H_5O_3$	H^+	$C_3H_5O_3^-$
I	.100	0	0
C	– x	x	x
E	.100 – x	x	x

Write the equation for K_a

$$K_a = \frac{[H^+]\,[C_3H_5O_3^-]}{[HC_3H_5O_3]}$$

Next find H^+

$$H^+ = 10^{-2.44} = .0036m = x$$

Therefore: $HC_3H_5O_3$ concentration at equilibrium

$$= (.100 - x) = .100 - .0036 = .096$$

$$K_a = \frac{[x]\,[x]}{[.100 - x]} = \frac{[.0036]\,[.0036]}{[.096]}$$

$$= 1.4 \times 10^{-4}$$

B. Calculate values of $[H^+]$ and pH of a solution of a weak acid/base.

The equation for a weak base is as follows. Use Kb, as OH^- is produced.

$$N_2H_4 + H_2O \rightarrow N_2H_5^+ + OH^-$$

$$Kb = \frac{[N_2H_5^+]\,[OH^-]}{[N_2H_4]}$$

Calculate the ph of a solution of a weak base, using the following equation:

$$N_2H_4 + H_2O \rightarrow N_2H_5^+ + OH^-$$

Original concentration of N_2H_4 = .100m

K_b= N_2H_4 = 1.3 x 10^{-6}

Use ICE: $N_2H_4 + H_2O \rightarrow N_2H_5^+ + OH^-$

N_2H_4 + $N_2H_5^+$ □ OH^-

	N_2H_4	**$N_2H_5^+$**	**OH^-**
I	.100	0	0
C	– x	x	x
E	.100–x	x	x

$$K_b = 1.3 \times 10^{-6} = \frac{(x)(x)}{(.100 - x)}$$

Since K_b is a very small number, ignore the –x component in the equilibrium.

($H_b \geq 400 \times K_b$; $.100 \geq 400 \times 1.3 \times 10^{-6}$)

For example: $1.3 \times 10^{-6} = \frac{(x)(x)}{.100}$

$1.3 \times 10^{-7} = x^2$

$3.61 \times 10^{-4} = x = OH^-$ concentration

$-\log [3.61 \times 10^{-4}] = pOH = 3.44$

pOH + pH = 14

pH = 14 – pOH = 14 – 3.44

pH = 10.56

C. <u>Calculate pH of salt solutions.</u>

Calculate salt solutions of weak acids and weak bases. When a salt is dissolved in H_2O, it will form a strong base (NaOH) and weak acid (HOCl).

$$NaOCl \rightarrow Na^+ + OCl^-$$

Does not impact because it is a spectator ion.

$OCl^- + H_2O \rightarrow HOCl + OH^-$

K_b is used because OH^- is generated. If H^+ is generated, use K_a.

$$K_b = \frac{[HOCl]\,[OH^-]}{[OCl^-]}$$

$NH_4Cl \rightarrow NH_4^+ + Cl^-$

$NH_4^+ + H_2O \rightarrow NH_3 + H_3O^+$ or H^+ (Depending on instructor or book provided)

$$K_b = \frac{[H^+]\,[NH_3]}{[NH_4^+]}$$

Example 1:

What is the ph of a 1.00 molar solution of NaCN? K_a for HCN = 6.2×10^{-10}

Solution Steps:

1. Note that this is a salt solution of a strong base (NAOH) and a weak acid (HCN) which will produce OH^- ions.
   $NaCN \rightarrow Na^+ + CN^-$
   $CN^- + H_2O \rightarrow HCN + OH^-$
2. Set up the ionization constant relationship:
   $$K_b = \frac{[HCN][OH^-]}{[CN^-]}$$
   (K_b is used when OH^- are produced; K_a is used when H^+ ions are produced.)
3. Find K_b

   $K_a \cdot K_b = 1 \times 10^{-14}$

   $$K_b = \frac{1 \times 10^{-14}}{K_a}$$

   $$K_b = \frac{1 \times 10^{-14}}{6.2 \times 10^{-10}}$$

   $K_b = 1.61 \times 10^{-5}$

4. Solve for $[OH^-]$ the known concentration of CN^- = 1 molar.
   $$1.61 \times 10^{-5} = \frac{[x][x]}{1}$$
   $1.61 \times 10^{-5} = x^2$
   $.004 = x = OH^-$ concentration
   pOH^- – Log $[OH^-]$ = 2.40

Find pH using the following formula:

$$pH + pOH = 14$$

$$pH = 14 - pOH$$

$$pH = 14 - 2.40$$

$$pH = 11.60$$

Example 2 Calculate the pH of .15 m CH_3NH_3Cl.

$K_b = 4.4 \times 10^{-4}$ (weak base/strong acid)

Equation: $CH_3NH_3Cl \rightarrow CH_3NH_3^+ + Cl^-$

$CH_3NH_3^+ + H_2O \rightarrow CH_3NH_2 + H_3O^+$ or H^+

$CH_3NH_3^+$ $CH_3NH_2 + H^+$

	$\underline{CH_3NH_3^+}$	$\underline{CH_3NH_2}$	+ $\underline{Cl^-}$
I	.15m	0	0
C	-x	x	x
E	.15–x	x	x

Determine if –x is significant. Apply test.

$$K_a = \frac{Kw}{K_b} = \frac{1.0 \times 10^{-14}}{4.4 \times 10^{-4}} = 2.3 \times 10^{-11}$$

$$K_a = \frac{[H^+][CH_3NH_2]}{CH_3NH_3^+} = 2.3 \times 10^{-11}$$

$$Ha \geq 400 \times K_a$$

$$.15 \geq 400 \times 2.3 \times 10^{-11}$$

Therefore, x is a very small amount compared to the original equation and can be ignored.

$$2.3 \times 10^{-11} = \frac{[x][x]}{.15}$$

$$1.9 \times 10^{-6} = x$$

$$pH = -\log[1.9 \times 10^{-6} m] = 5.72$$