

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:



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

Perform ICE:	<u>$HC_3H_5O_3$</u>	<u>H^+</u>	<u>$C_3H_5O_3^-$</u>
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

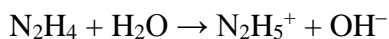
$$= (.100 - x) = .100 - .0036 = .096$$

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

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

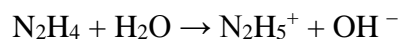
B. Calculate values of $[H^+]$ and pH of a solution of a weak acid/base.

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



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

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



Original concentration of $\text{N}_2\text{H}_4 = .100\text{M}$

$$K_b = \text{N}_2\text{H}_4 = 1.3 \times 10^{-6}$$

Use ICE: $\text{N}_2\text{H}_4 + \text{H}_2\text{O} \rightarrow \text{N}_2\text{H}_5^+ + \text{OH}^-$



	<u>N_2H_4</u>	<u>N_2H_5^+</u>	<u>OH^-</u>
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.

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

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

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

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

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

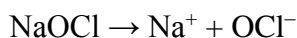
$$\text{pOH} + \text{pH} = 14$$

$$\text{pH} = 14 - \text{pOH} = 14 - 3.44$$

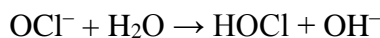
$$\text{pH} = 10.56$$

C. Calculate pH of salt solutions.

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).

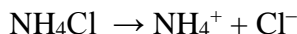


↑
→ Does not impact because it is a spectator ion.



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

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



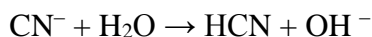
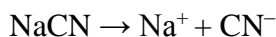
$$K_b = \frac{[\text{H}^+][\text{NH}_3]}{[\text{NH}_4^+]}$$

Example 1:

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

Solution Steps:

- Note that this is a salt solution of a strong base (NaOH) and a weak acid (HCN) which will produce OH^- ions.



- Set up the ionization constant relationship:

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

(K_b is used when OH^- are produced; K_a is used when H^+ ions are produced.)

- Find K_b

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

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

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

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

- Solve for $[\text{OH}^-]$ the known concentration of $\text{CN}^- = 1$ molar.

$$1.61 \times 10^{-5} = \frac{[x][x]}{1}$$

$$1.61 \times 10^{-5} = x^2$$

$$.004 = x = \text{OH}^- \text{ concentration}$$

$$\text{pOH}^- - \text{Log} [\text{OH}^-] = 2.40$$

Find pH using the following formula:

$$\text{pH} + \text{pOH} = 14$$

$$\text{pH} = 14 - \text{pOH}$$

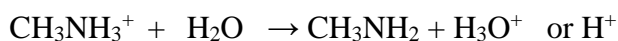
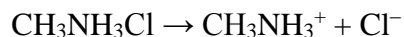
$$\text{pH} = 14 - 2.40$$

$$\text{pH} = 11.60$$

Example 2 Calculate the pH of .15 m $\text{CH}_3\text{NH}_3\text{Cl}$.

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

Equation:



I	.15m	0	0
C	-x	x	x
E	.15-x	x	x

Determine if $-x$ is significant. Apply test.

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

$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{NH}_2]}{\text{CH}_3\text{NH}_3^+} = 2.3 \times 10^{-11}$$

$$\text{Ha} \geq 400 \times K_a$$

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

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

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

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

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