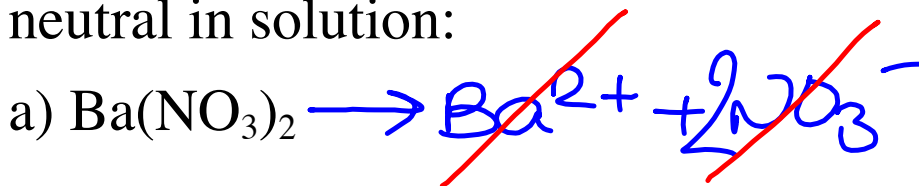
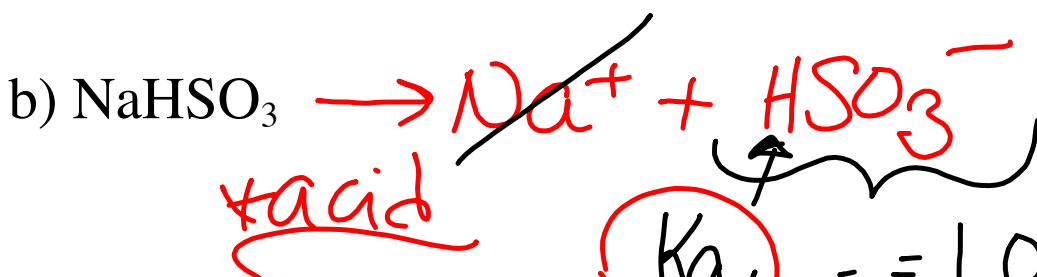


Predict whether the following are acidic, basic or neutral in solution:



neutral



acid

$$K_b = \frac{K_w}{K_a(\text{H}_2\text{SO}_3)} = \frac{1 \times 10^{-14}}{6.5 \times 10^{-2}} = 1.5 \times 10^{-13}$$

$K_a(\text{HSO}_3^-) = 1.0 \times 10^{-7}$



Basic

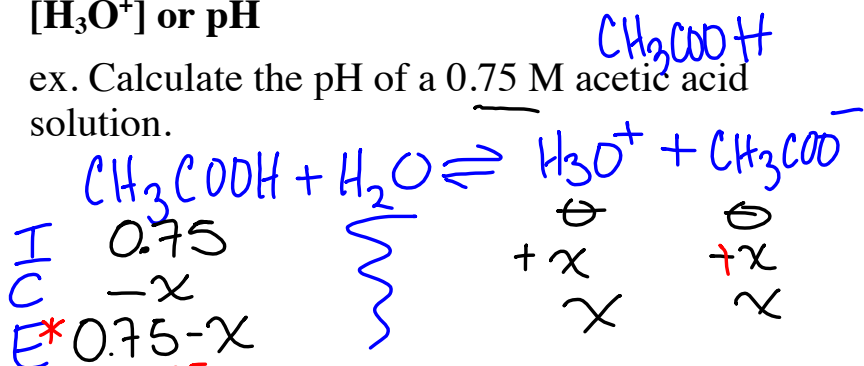
## Calculations involving Ka & Kb

### 1. Calculations involving Ka:

- when a weak acid (HA) is put into water, some of the acid ionizes
- therefore, a certain amount of  $\text{H}_3\text{O}^+$  is produced
- the smaller the  $K_a$ , the less  $\text{H}_3\text{O}^+$  produced
- many of these problems require an ICE table

#### A - Using initial [acid] & $K_a$ to determine $[\text{H}_3\text{O}^+]$ or pH

ex. Calculate the pH of a 0.75 M acetic acid solution.



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = 1.8 \times 10^{-5}$$

*Handwritten notes:*  $\approx 0.75$  written below  $0.75-x$ .  $0.00018$  written in red above the  $1.8 \times 10^{-5}$ .

*Handwritten notes:* assume  $0.75-x \approx 0.75$   
\*can only make assumption if  $[ ] > 1000x$

$$K_a = \frac{x^2}{0.75} = 1.8 \times 10^{-5}$$

$$x^2 = (0.75)(1.8 \times 10^{-5})$$

$$= \sqrt{(0.75)(1.8 \times 10^{-5})}$$

$$= 0.0036743$$

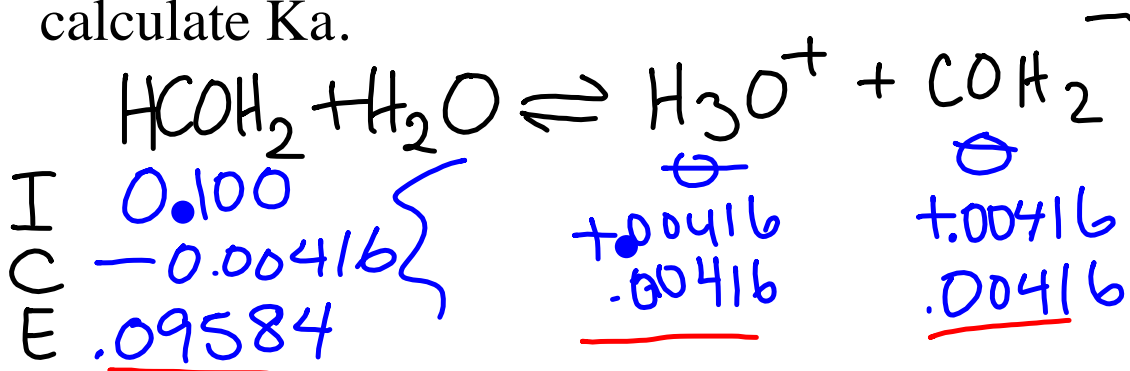
$$= [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] \\ = -\log(0.0036743)$$

$$= \underline{2.43}$$

## B - Calculating $K_a$ from $[H_3O^+]$ or pH.

ex. If the pH of 0.100 M  $HCOH_2$  is 2.38 at  $25^\circ C$ , calculate  $K_a$ .



$$K_a = \frac{[H_3O^+][COH_2^-]}{[HCOH_2]} = ? = \frac{(.00416)(.00416)}{.09584}$$

$$pH = 2.38$$

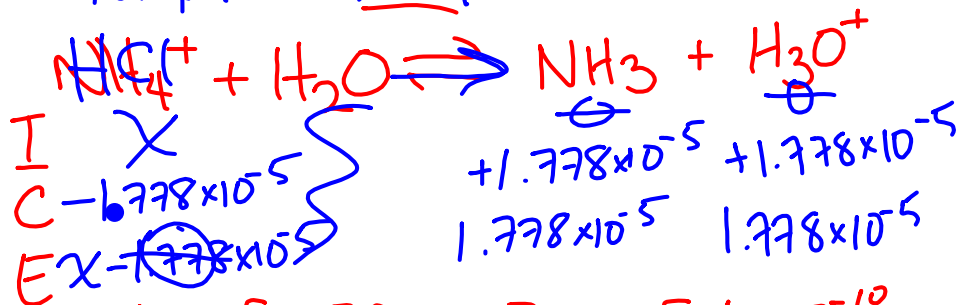
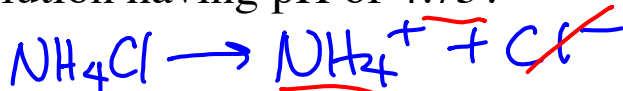
$$[H_3O^+] = \text{antilog } -2.38$$

eg = .0041686

$$= 1.806 \times 10^{-4}$$
$$1.8 \times 10^{-4}$$

## C - Finding the initial concentration of a weak acid.

ex. What mass of  $\text{NH}_4\text{Cl}$  will produce 1.50 L of a solution having pH of 4.75?



$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = 5.6 \times 10^{-10}$$

pH = 4.75

$[\text{H}_3\text{O}^+] = \text{antilog}^{-4.75}$   
 $= 1.778 \times 10^{-5} \text{ M}$

$$K_a = \frac{(1.778 \times 10^{-5})^2}{x - 1.778 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$$\frac{(1.778 \times 10^{-5})^2}{5.6 \times 10^{-10}} = x - 1.778 \times 10^{-5}$$

$$x = \left[ \frac{(1.778 \times 10^{-5})^2}{5.6 \times 10^{-10}} \right] + 1.778 \times 10^{-5}$$

$$x = 0.5645 \text{ M} = [\text{NH}_4^+] = [\text{NH}_4\text{Cl}]$$

$$g_{\text{NH}_4\text{Cl}} = \left( \frac{0.5645 \text{ mol}}{\text{L}} \right) (1.5 \text{ L}) \left( \frac{53.5 \text{ g}}{\text{mol}} \right)$$

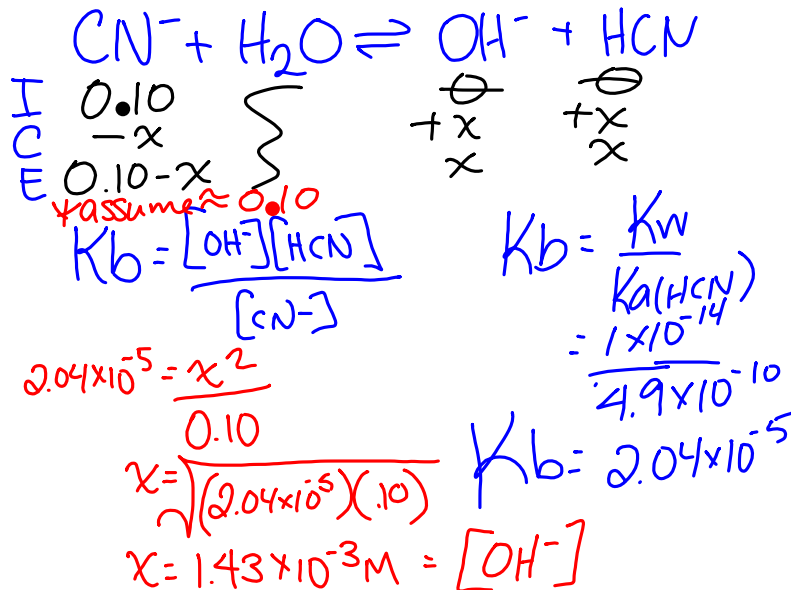
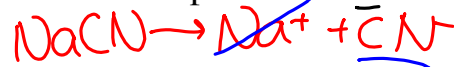
$$= 45 \text{ g}$$

## 2. Calculations involving $K_b$

- calculations involving weak bases are similar to the calculations involving weak acids with two important differences:
  - the  $K_b$  value must be calculated
  - the resulting solution will be basic, not acidic, which means using the  $[OH^-]$  rather than  $[H_3O^+]$

### A - Using initial concentration of base and $K_b$ to determine pH

ex. Calculate the pH of a 0.10 M NaCN solution.



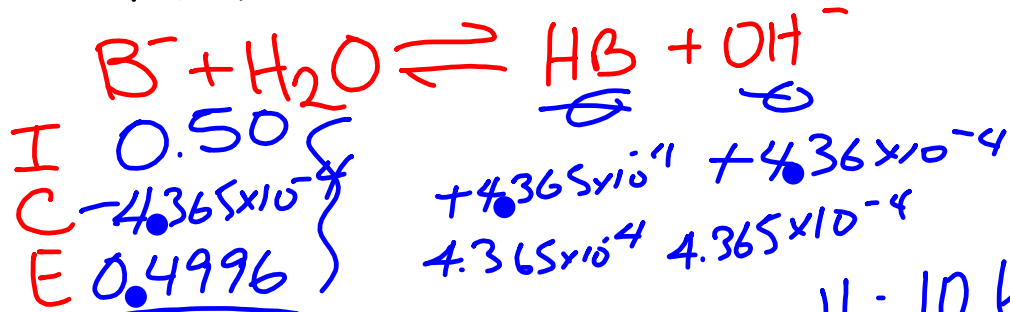
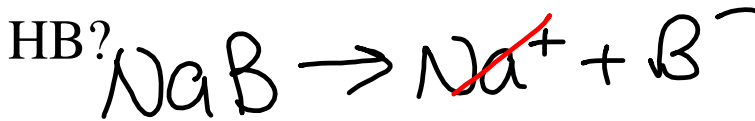
$$pOH = -\log[OH^-]$$
$$= -\log(1.43 \times 10^{-3})$$

$$= 2.845$$
$$pH = 14 - pOH$$
$$= 14 - 2.845$$

$$pH = 11.15$$

## B - Calculating Kb from pOH or pH

ex. The pH of a 0.50 M solution of the weak base NaB is 10.64. What is Ka for the conjugate acid HB?



$$K_b = \frac{[\text{HB}][\text{OH}^-]}{[\text{B}^-]}$$

$$K_b = \frac{(4.365 \times 10^{-4})^2}{0.4996}$$

$$K_b = 3.8139 \times 10^{-7}$$

$$K_a = \frac{K_w}{K_b} = \frac{1 \times 10^{-14}}{3.8139 \times 10^{-7}}$$

$$= 2.6219 \times 10^{-8}$$

$$2.6 \times 10^{-8}$$

Ka

$$\text{pH} = 10.64$$

$$\text{pOH} = 14 - 10.64$$

$$\text{pOH} = 3.36$$

$$[\text{OH}^-] = \text{antilog}^{-3.36}$$

$$= 4.365 \times 10^{-4}$$