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

a) $Ba(NO_3)_2 \longrightarrow \cancel{Ba^{2+}} + \cancel{2NO_3^-}$

neutral

b) $NaHSO_3 \longrightarrow \cancel{Na^+} + \underbrace{HSO_3^-}$

acid

$K_{a(HSO_3^-)} = 1.0 \times 10^{-7}$

$K_b = \frac{K_w}{K_{a(H_2SO_3)}} = \frac{1 \times 10^{-14}}{1.5 \times 10^{-2}} = 6.7 \times 10^{-13}$

c) $Na_3PO_4 \rightarrow \cancel{3Na^+} + PO_4^{3-}$

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 H_3O^+ is produced
- the smaller the Ka, the less H_3O^+ produced
- many of these problems require an ICE table

A - Using initial [acid] & Ka to determine $[H_3O^+]$ or pH

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

CH_3COOH

$$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$$

	CH_3COOH	H_2O	H_3O^+	CH_3COO^-
I	0.75		0	0
C	$-x$		$+x$	$+x$
E*	$0.75-x$ ≈0.75		x	x

$$Ka = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = 1.8 \times 10^{-5}$$

0.00018

*assume $0.75 - x \approx 0.75$

*can only make assumption if [ ] > 1000x K

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

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

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

$$= 0.0036743 \text{ M}$$

$$= [H_3O^+]$$

$$pH = -\log[H_3O^+]$$

$$= -\log(.0036743)$$

$$= \underline{2.43}$$

B - Calculating Ka from [H_3O^+] or pH.

ex. If the pH of 0.100 M $HCOH_2$ is 2.38 at 25°C, calculate Ka.

$$HCOH_2 + H_2O \rightleftharpoons H_3O^+ + COH_2^-$$

	$HCOH_2$	H_3O^+	COH_2^-
I	0.100	0	0
C	−0.00416	+.00416	+.00416
E	.09584	.00416	.00416

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

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

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

pH = 2.38

$[H_3O^+]_{eq}$ = antilog −2.38

= .0041686

C - Finding the initial concentration of a weak acid.

ex. What mass of $\underbrace{NH_4Cl}_{Salt}$ will produce 1.50 L of a solution having pH of 4.75?

$NH_4Cl \rightarrow NH_4^+ + \cancel{Cl^-}$

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$

I x 0   0

C -1.778×10^{-5} $+1.778\times10^{-5}$ $+1.778\times10^{-5}$

E $x-1.778\times10^{-5}$ 1.778×10^{-5} 1.778×10^{-5}

$Ka = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.6\times10^{-10}$

pH = 4.75

$[H_3O^+] = \text{antilog} -4.75$

EQ $= 1.778\times10^{-5}$ M

$Ka = \frac{(1.778\times10^{-5})^2}{x-1.778\times10^{-5}} = 5.6\times10^{-10}$

$\frac{(1.778\times10^{-5})^2}{5.6\times10^{-10}} = x - 1.778\times10^{-5}$

$x = \left[\frac{(1.778\times10^{-5})^2}{5.6\times10^{-10}}\right] + 1.778\times10^{-5}$

$x = 0.5645\,M = [NH_4^+] = [NH_4Cl]$

$g_{NH_4Cl} = \left(\frac{0.5645\,mol}{\cancel{L}}\right)(1.5\,\cancel{L})\left(\frac{53.5\,g}{mol}\right)$

$= 45\,g$

2. Calculations involving Kb

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

A - Using initial concentration of base and Kb to determine pH

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

$NaCN \rightarrow \cancel{Na^+} + CN^-$

$CN^- + H_2O \rightleftharpoons OH^- + HCN$

	CN^-	H_2O	OH^-	HCN
I	0.10		0	0
C	$-x$		$+x$	$+x$
E	$0.10-x$		x	x

*assume ≈ 0.10

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

$Kb = \frac{Kw}{Ka(HCN)} = \frac{1\times10^{-14}}{4.9\times10^{-10}}$

$Kb = 2.04\times10^{-5}$

$2.04\times10^{-5} = \frac{x^2}{0.10}$

$x = \sqrt{(2.04\times10^{-5})(.10)}$

$x = 1.43\times10^{-3}M = [OH^-]$

$pOH = -\log[OH^-]$
$= -\log(1.43\times10^{-3})$
$= 2.845$

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

$\boxed{pH = 11.15}$

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?

$NaB \rightarrow \cancel{Na}^+ + B^-$

$B^- + H_2O \rightleftharpoons HB + OH^-$

	B^-	HB	OH^-
I	0.50	0	0
C	-4.365×10^{-4}	$+4.365\times10^{-4}$	$+4.365\times10^{-4}$
E	0.4996	4.365×10^{-4}	4.365×10^{-4}

pH = 10.64

pOH = 14 - 10.64

pOH = 3.36

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

$= 4.365\times10^{-4}$

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

$$Kb = \frac{(4.365\times10^{-4})^2}{.4996}$$

$$Kb = 3.8139\times10^{-7}$$

$$Ka = \frac{Kw}{Kb} = \frac{1\times10^{-14}}{3.8139\times10^{-7}} = 2.6219\times10^{-8}$$

2.6×10^{-8}

Ka