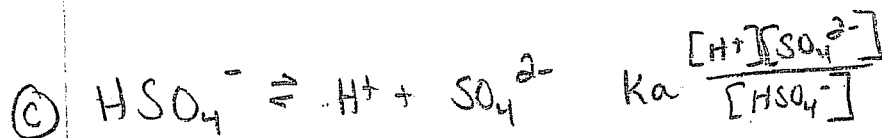
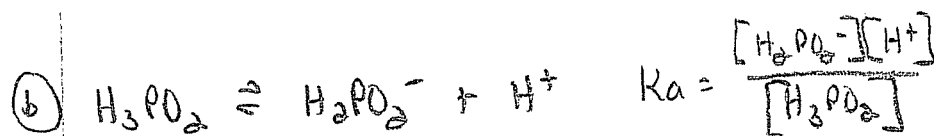
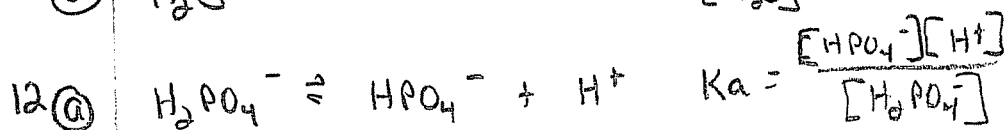
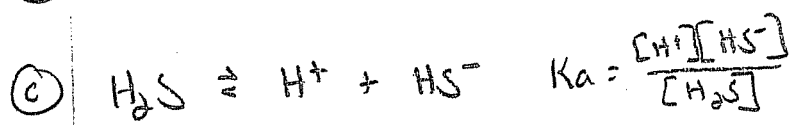
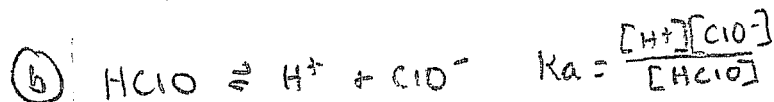
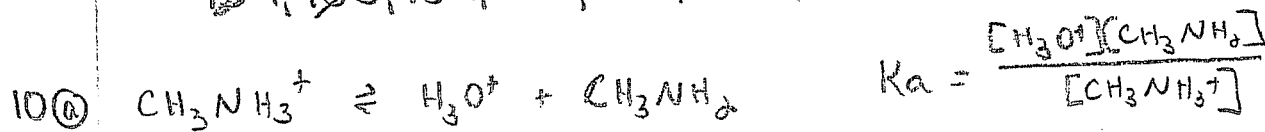


Ch. 18 HW ~~10, 12, 14, 16, 18, 24, 26, 28, 30, 32, 34, 35,~~
~~42, 44, 46, 48, 52, 54, 56, 58, 60, 62, 64, 66, 68, 70, 72, 74, 76,~~
~~78, 80, 82, 84, 86, 88, 90, 92, 94, 96, 100, 102, 104, 106, 108,~~
~~110, 112, 114, 116, 120, 122, 124, 126, 128,~~
~~130, 132, 134, 136, 138, 140, 142, 144, 146, 148, 150,~~
~~152, 154, 156, 158, 160, 162, 164, 166, 168, 170, 172,~~
~~174, 176, 178, 180~~



14 Decreasing Acid Strength
 $\text{HCl} > \text{HNO}_2 > \text{HClO} > \text{HCN}$
 $\text{HClO} = 2.9 \times 10^{-8}$
 $\text{HCN} = 6.2 \times 10^{-10}$
 $\text{HNO}_2 = 7.1 \times 10^{-4}$

- 16 a) CH_3NH_2 weak base
 b) K_2O strong base - why? water soluble containing O^{2-}
 c) HI strong acid
 d) HCOOH weak acid

18a) $\text{HOCH}_2\text{CH}_2\text{NH}_2$ amine weak base

b) H_2SeO_4 strong acid

c) HS^- weak acid

d) B(OH)_3 why weak acid?

Alternate formula is $\text{H}_3\text{BO}_3 \text{H}-\text{O}-\text{B}(\text{O}-\text{H})_2$

willlose H^+ ions

24 a) pH of 0.0333 M HNO_3

b) pOH of 0.0347 KOH

$-\log(0.033) = 1.48$ acidic
 $-\log(0.0347) = 1.46$ pOH = 12.5 basic

26 a) pH of 7.52×10^{-4} CsOH?

pOH = $-\log(7.52 \times 10^{-4})$

pOH = 3.12

pH = 10.87 basic

b) 1.59×10^{-3} M HClO_4

pH = 2.8

pOH = 11.20

28 a) pH = 3.47

pOH = 10.53

$[\text{H}^+] = 10^{-3.47} = 3.39 \times 10^{-4}$ M

$[\text{OH}^-] = 10^{-10.53} = 2.95 \times 10^{-11}$ M

b) pOH = 4.33

pH = 9.67

$[\text{H}^+] = 10^{-9.67} = 2.14 \times 10^{-10}$ M

$[\text{OH}^-] = 10^{-4.33} = 4.68 \times 10^{-5}$ M

30 a) pH = 8.97

pOH = 5.03

$[\text{H}^+] = 1.07 \times 10^{-9}$ M

$[\text{OH}^-] = 9.33 \times 10^{-6}$ M

b) pOH = 11.27

pH = 2.73

$[\text{OH}^-] = 5.37 \times 10^{-12}$ M

$[\text{H}^+] = 1.86 \times 10^{-3}$ M

32) pH = 9.33 \rightarrow pH = 9.07

$[\text{H}^+] = 4.677 \times 10^{-10}$ M

$[\text{H}^+] = 8.51 \times 10^{-10}$ M

add 3.83×10^{-10} mol H^+

34) 87.5 mL of HA to adjust pH = 8.92 to pH = 6.33

pH = 8.92

$[\text{H}^+] = 1.20 \times 10^{-9}$ M

pH = 6.33

$[\text{H}^+] = 4.68 \times 10^{-7}$ M

add 4.66×10^{-7} mol/L if making 1 L

So $\frac{4.66 \times 10^{-7} \text{ mol}}{1 \text{ L}} = \frac{x}{0.0875}$

$x = 4.078 \times 10^{-8}$ mol needed

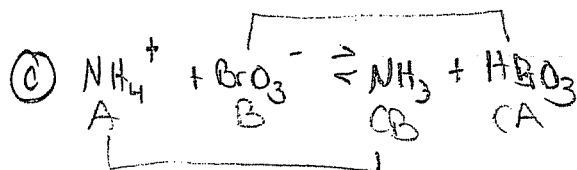
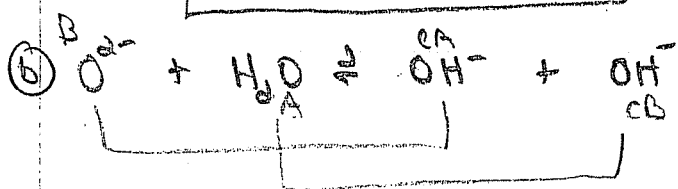
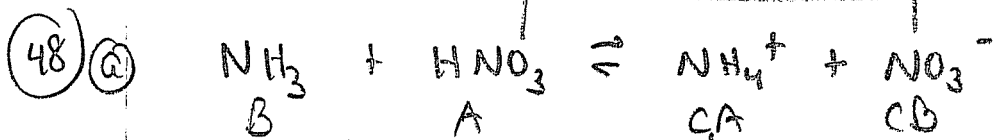
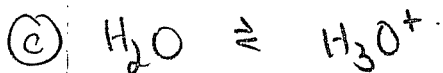
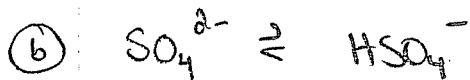
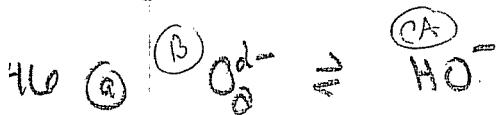
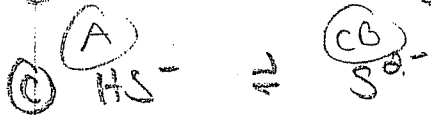
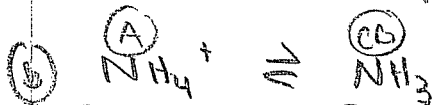
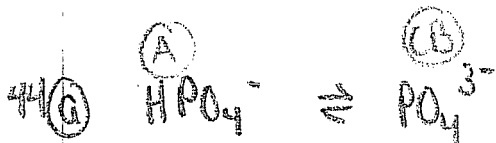
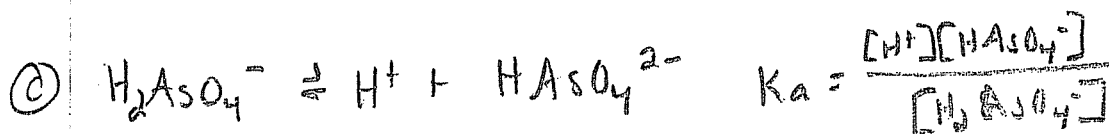
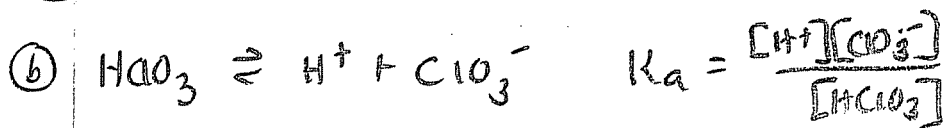
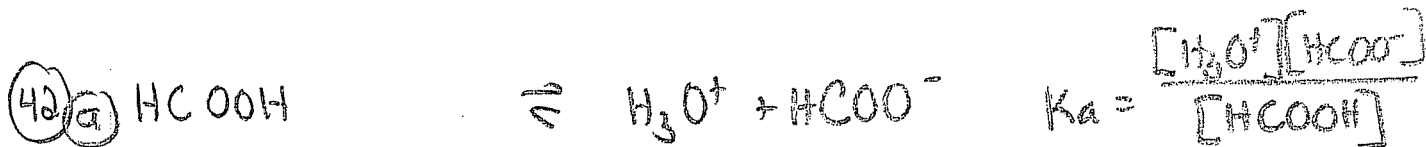
35

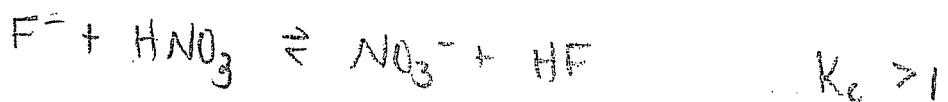
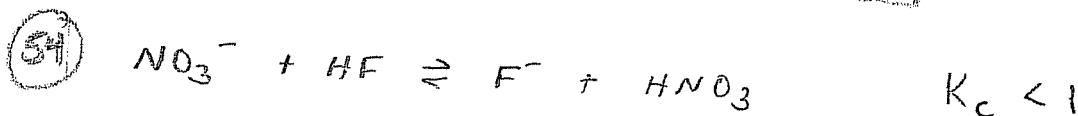
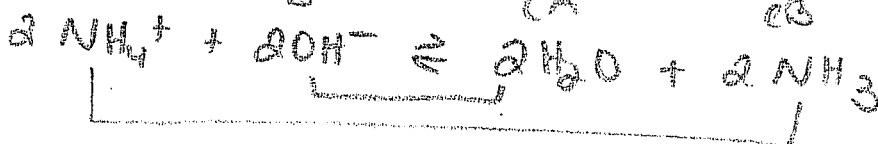
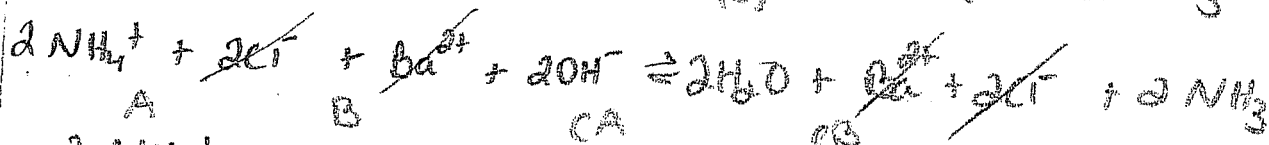
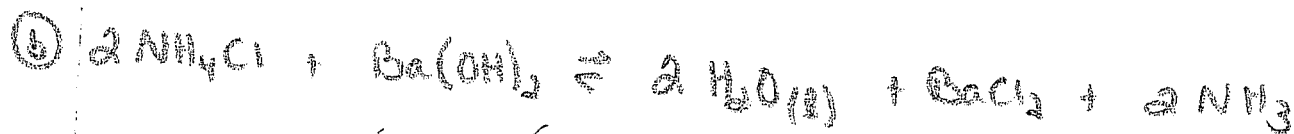
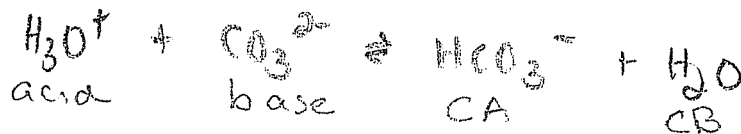
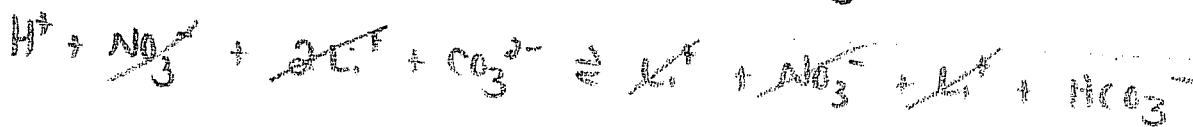
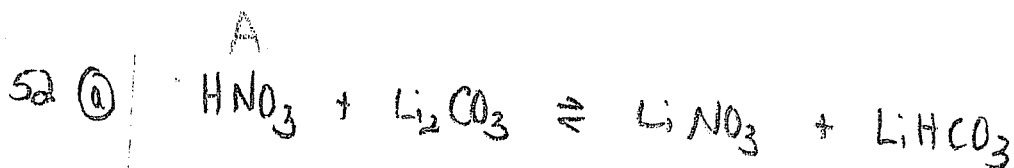
H_3O^+ = purple
Volume = equal

$pH_A = 4.8$
 $pH_B =$

$[H^+] = 10^{-4.8} = 1.58 \times 10^{-5} M$
 $[H^+] =$

2 molec.
25 molec.




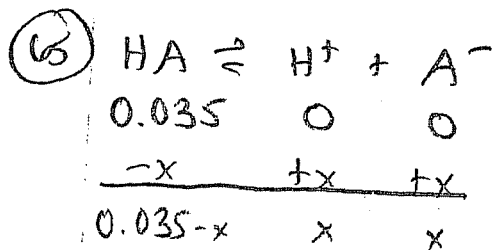


(61) 0.15M HA blue+green 33% dissociated

(1) 5 HA
4 dissociated
 $\frac{4}{9} \times 100 = 44\%$

(2) 7 HA
2 dissociated
 $\frac{2}{9} \times 100 = 22\%$

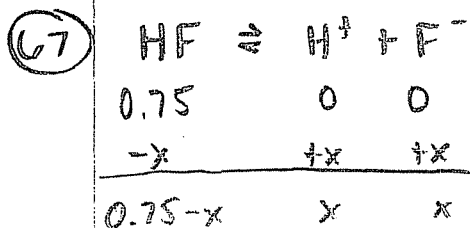
(3)  6 HA
3 dissociated
 $\frac{3}{9} = 33\%$



pH = 4.88
 $[\text{H}^+] = 10^{-4.88}$
 $= 1.32 \times 10^{-5}$

$$K_a = \frac{(1.32 \times 10^{-5})^2}{0.035}$$

$$K_a = 4.97 \times 10^{-9}$$

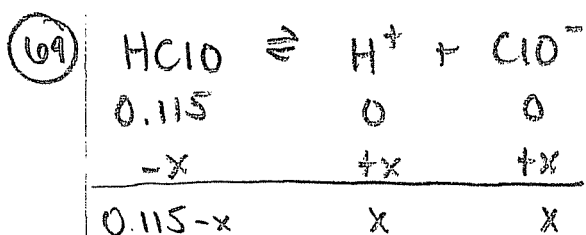


$$K_a = 6.8 \times 10^{-4}$$

$$6.8 \times 10^{-4} = \frac{x^2}{0.75}$$

$$x = 0.0226 \text{ M} = [\text{H}^+] = [\text{F}^-]$$

$$[\text{OH}^-] = \frac{1 \times 10^{-14}}{0.0226} = 4.4 \times 10^{-13} \text{ M}$$



$$pK_a = 7.54$$

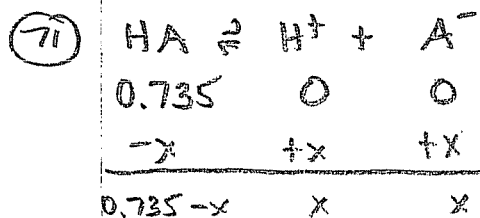
$$K_a = 10^{-7.54} = 2.88 \times 10^{-4}$$

$$2.88 \times 10^{-4} = \frac{x^2}{0.115}$$

$$x = 5.76 \times 10^{-5} \text{ M} = [\text{H}^+] = [\text{ClO}^-]$$

$$pH = 4.24 \checkmark$$

$$[\text{HClO}] = 0.115 - 5.76 \times 10^{-5} = 0.115 \text{ M}$$



$$\% \text{ dissociated} = 12.5\%$$

$$0.125 = \frac{x}{0.735}$$

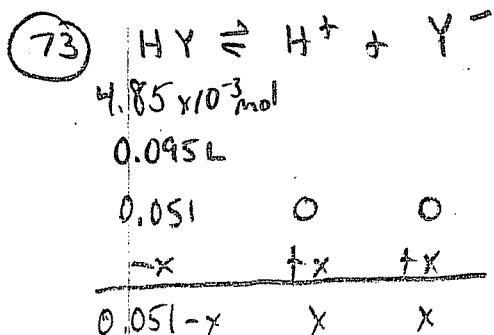
$$x = 0.0919 \text{ M} = [\text{H}^+] = [\text{A}^-]$$

$$pH = 1.04$$

$$pOH = 12.96$$

$$[\text{OH}^-] = 1.096 \times 10^{-13} \text{ M}$$

$$K_a = \frac{(0.0919)^2}{0.735} = 0.0115$$



$$pH = 2.68 \quad K_a = ?$$

$$[\text{H}^+] = 2.09 \times 10^{-3} \text{ M}$$

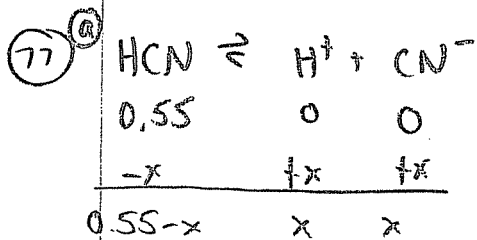
$$K_a = \frac{(2.09 \times 10^{-3})^2}{0.051}$$

$$K_a = 8.56 \times 10^{-5}$$

If subtract x

$$K_a = \frac{(2.09 \times 10^{-3})^2}{0.051 - 2.09 \times 10^{-3}}$$

$$K_a = 8.93 \times 10^{-5}$$



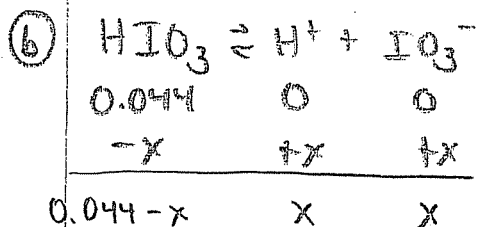
$$K_a = 6.2 \times 10^{-10}$$

$$6.2 \times 10^{-10} = \frac{x^2}{0.55}$$

pH = ?

$$x = 1.85 \times 10^{-5} = [\text{H}^+]$$

$$\text{pH} = 4.73$$



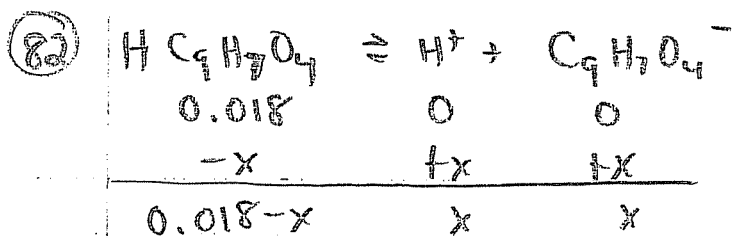
$$K_a = 0.016$$

$$0.016 = \frac{x^2}{0.044}$$

$$x = 0.0265$$

$$\text{pH} = 1.58 \quad \text{pOH} = 12.42$$

← technically should not neglect x & use quadratic



$$K_a = 3.6 \times 10^{-4}$$

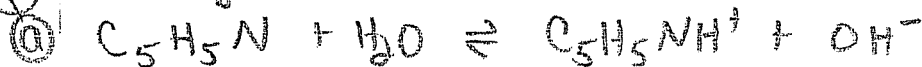
$$3.6 \times 10^{-4} = \frac{x^2}{0.018}$$

$$x = 0.00255$$

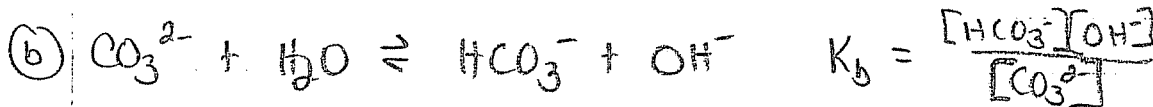
$$\text{pH} = 2.59$$

99

Write K_b



$$K_b = \frac{[\text{C}_5\text{H}_5\text{NH}^+][\text{OH}^-]}{[\text{C}_5\text{H}_5\text{N}]}$$



89 Choose weaker acid

a) HBr or H_2Se

★ weaker because less EN = stronger bond = less dissociation = weaker acid

b) HClO_4 ↑ more EN
 H_2SO_4 ↑ weaker acid

c) H_2SO_3 ↑ weaker less oxygens
 H_2SO_4

91(a) HI or HBr

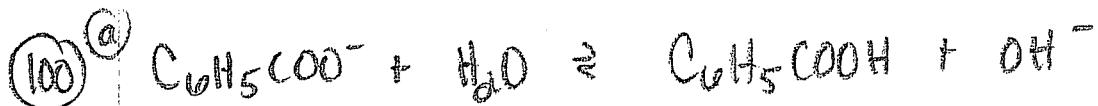
↑
Smaller radius = strong bond
= weak acid

(b) H_3AsO_4 H_2SeO_4

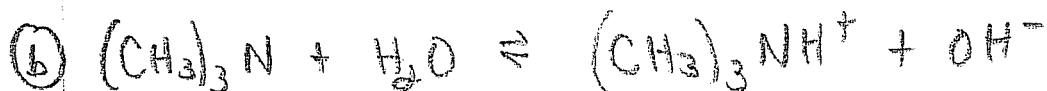
↑
weaker
less EN

(c) HNO_3 HNO_2

↓ less # oxygens = stronger bond = weaker acid



$$K_b = \frac{[C_6H_5COOH][OH^-]}{[C_6H_5COO^-]}$$



$$K_b = \frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]}$$

(104) pH of dimethylamine



$K_b = 5.9 \times 10^{-4}$

0.12 -x	0 +x	0 +x
0.12-x	x	x

$$5.9 \times 10^{-4} = \frac{x^2}{0.12}$$

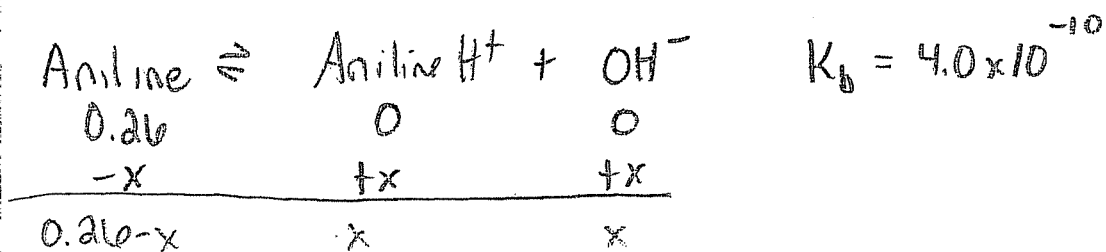
$$x = 0.008414$$

$$pOH = 2.07$$

$$pH = 11.93$$

Should use quadratic
but...
||
h

(106)



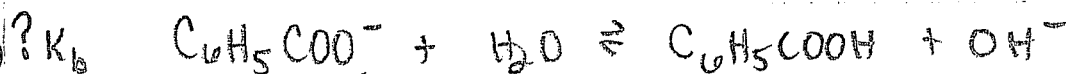
$$4.0 \times 10^{-10} = \frac{x^2}{0.26}$$

$$x = 1.02 \times 10^{-5}$$

$$\text{pOH} = 4.99$$

$$\text{pH} = 9.01$$

(108) (a)



$$K_a = 6.3 \times 10^{-5}$$

benzoic acid

$$K_a K_b = K_w$$

$$(6.3 \times 10^{-5})(K_b) = 1 \times 10^{-14}$$

$$K_b = 1.59 \times 10^{-10}$$



$$\text{p}K_b = 4.49$$

$$K_b = 3.24 \times 10^{-5}$$

$$K_a K_b = K_w$$

$$K_a (3.24 \times 10^{-5}) = 1 \times 10^{-14}$$

$$K_a = 3.09 \times 10^{-10}$$

(112) (a)



$$0.100$$

$$-x$$

$$0.100-x$$

$$0$$

$$+x$$

$$x$$

$$0$$

$$+x$$

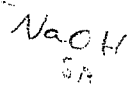
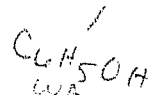
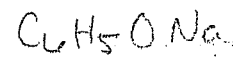
$$x$$

$$1 \times 10^{-4} = \frac{x^2}{0.1}$$

$$x = 0.00316 \text{ M}$$

$$\text{pOH} = 2.5$$

$$\text{pH} = 11.5$$

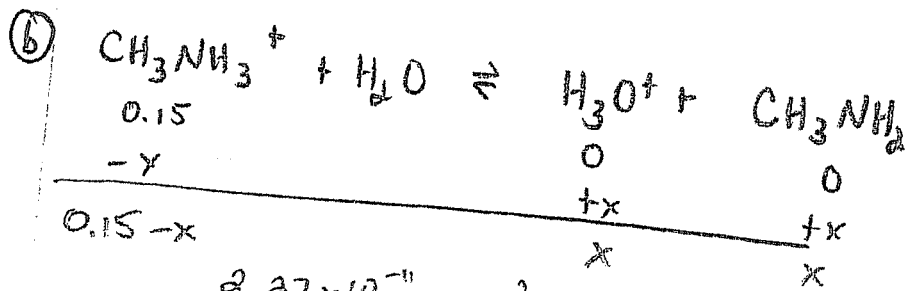


$$K_a = 1 \times 10^{-10}$$

$$K_b K_a = K_w$$

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

$$K_b = 1 \times 10^{-4}$$

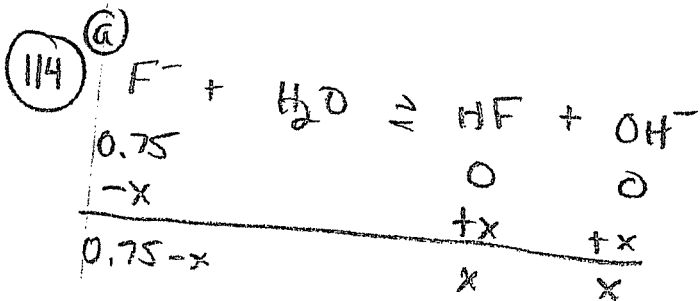


$K_b = 4.4 \times 10^{-4}$
 $K_a = 2.27 \times 10^{-11}$

$$2.27 \times 10^{-11} = \frac{x^2}{0.15}$$

$$x = 1.85 \times 10^{-6}$$

$$\text{pH} = 5.73$$

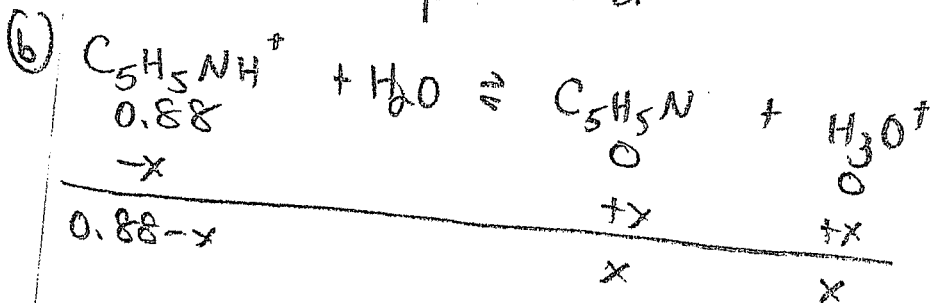


$K_a = 6.8 \times 10^{-4}$
 $K_b = 1.47 \times 10^{-11}$

$$1.47 \times 10^{-11} = \frac{x^2}{0.75-x}$$

$$3.32 \times 10^{-6} = x$$

$\text{pOH} = 5.48$
 $\text{pH} = 8.52$



$K_b = 1.7 \times 10^{-9}$
 $K_a = 5.88 \times 10^{-6}$

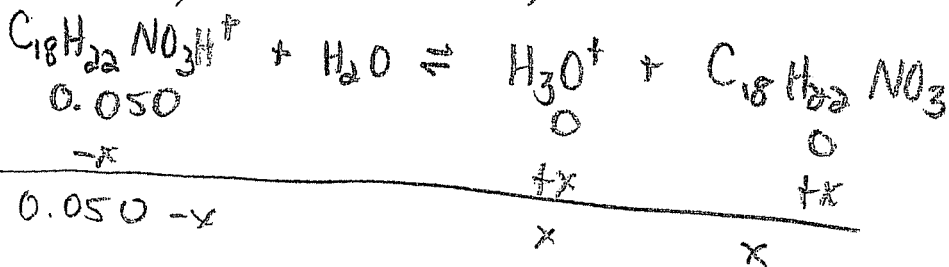
$$5.88 \times 10^{-6} = \frac{x^2}{0.88}$$

$$x = 0.002275$$

$$\text{pH} = 2.64$$

weak base
makes
acidic
solution

116

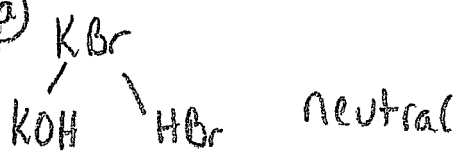


$pK_b = 5.80$
 $K_b = 1.58 \times 10^{-6}$
 $K_a = 6.3 \times 10^{-9}$

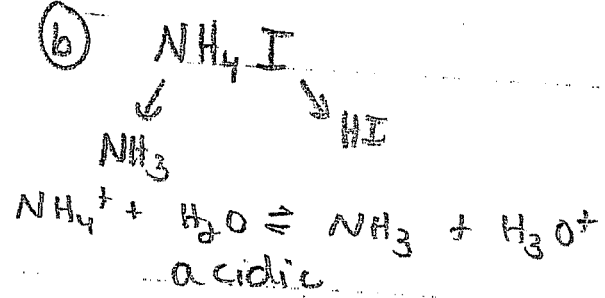
$$\begin{aligned}
 6.31 \times 10^{-9} &= \frac{x^2}{0.050} \\
 x &= 1.78 \times 10^{-5} \\
 \text{pH} &= 4.75
 \end{aligned}$$

120

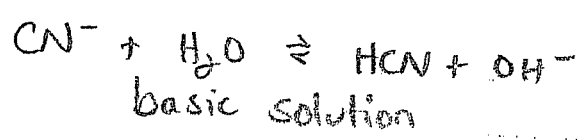
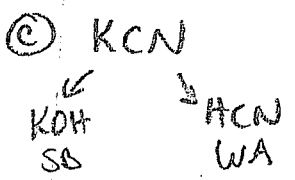
a



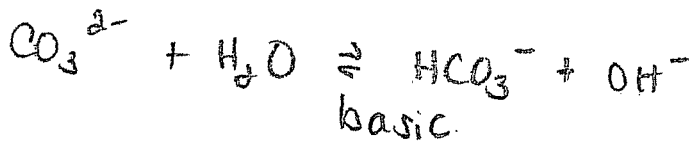
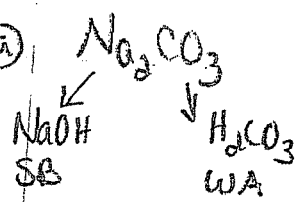
b



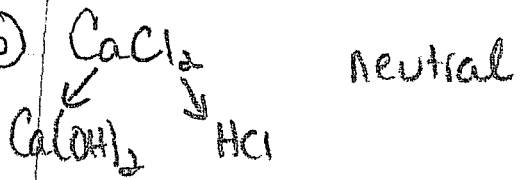
c



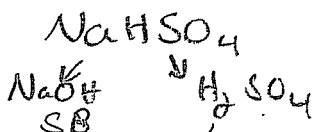
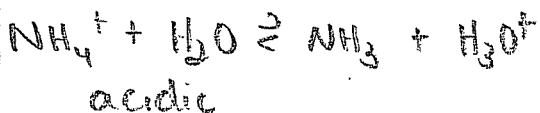
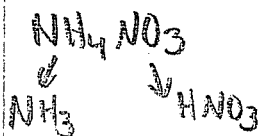
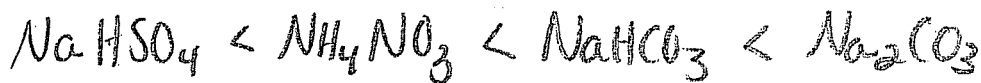
122a



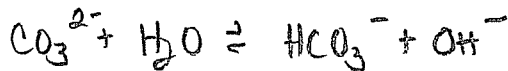
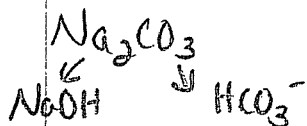
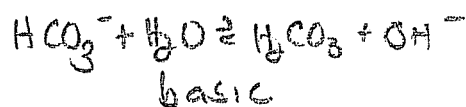
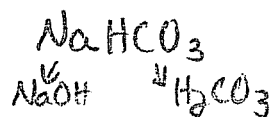
b



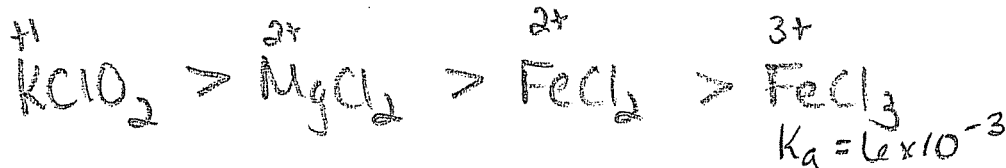
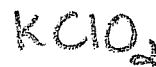
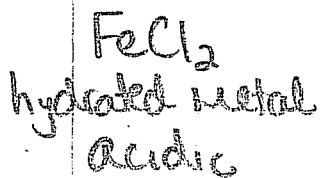
128 ⑥ ↑ pH



amphiprotic
in water



1296 decreasing pH



Smaller
+ more
charged =
more acidic

* K_a values would be helpful

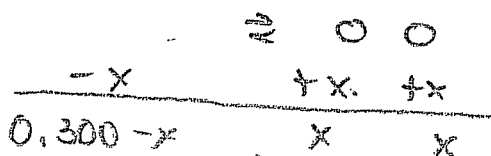


a) @ equilibrium

$$K = \frac{(4)^2}{8}$$

$$K = 2$$

b)



$$Q = \frac{2(2)}{6}$$

$$Q = 0.667$$

c) $Q = \frac{2(2)}{4}$
 $= 1$

d) $Q = \frac{2(2)}{2}$ ☆ at equilibrium
 $Q = 2$

155) a) higher HCl $\rightarrow H^+ + Cl^-$ Completely dissociates
 so produces more ions, higher electrical conductivity than weak acid

b) about the same because very similar to $[H_3O^+][OH^-]$
 is just plain water
 won't significantly change amt H^+ or OH^-