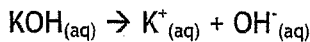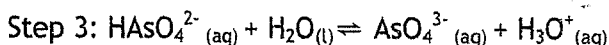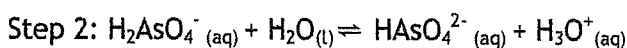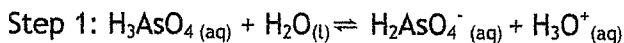# Acid-Base Equilibrium Problems

1. Write equations which represent the dissociation of each of these acids or bases in aqueous solution. Use a single arrow in the case of a strong acid or base, and a double arrow to represent the equilibrium condition that exists in the solution of a weak acid or base. Show each step of dissociation for polyprotic acids.
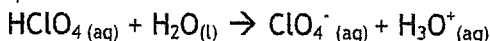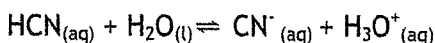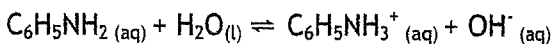
   a) KOH

   $$KOH_{(aq)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$$

   b) H_3AsO_4

   Step 1: $H_3AsO_{4\,(aq)} + H_2O_{(l)} \rightleftharpoons H_2AsO_4^-{}_{(aq)} + H_3O^+_{(aq)}$

   Step 2: $H_2AsO_4^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HAsO_4^{2-}{}_{(aq)} + H_3O^+_{(aq)}$

   Step 3: $HAsO_4^{2-}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons AsO_4^{3-}{}_{(aq)} + H_3O^+_{(aq)}$

   c) $HClO_4$

   $$HClO_{4\,(aq)} + H_2O_{(l)} \rightarrow ClO_4^-{}_{(aq)} + H_3O^+_{(aq)}$$

   d) $HCN_{(aq)}$

   $$HCN_{(aq)} + H_2O_{(l)} \rightleftharpoons CN^-_{(aq)} + H_3O^+_{(aq)}$$

   e) $C_6H_5NH_2$ (a weak base)

   $$C_6H_5NH_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons C_6H_5NH_3^+{}_{(aq)} + OH^-_{(aq)}$$

2. Benzoic acid, $HC_6H_5CO_2$, is an organic acid whose sodium salt, $NaC_6H_5CO_2$, has long been used as a safe food additive to protect beverages and many foods against harmful yeasts and bacteria. The acid is monoprotic. Write the K_a expression for this acid.

   $$HC_6H_5CO_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons C_6H_5CO_2^-{}_{(aq)} + H_3O^+_{(aq)}$$
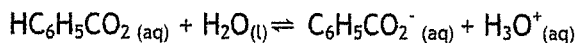
   $$K_a = \frac{[C_6H_6CO_2^-][H_3O^+]}{[HC_6H_6CO_2]}$$

3. Write the equilibrium equations and the K_b expressions for each of the following bases.

   a) CN^- (cyanide ion)

   $$CN^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HCN_{(aq)} + OH^-_{(aq)}$$

   $$K_b = \frac{[HCN][OH^-]}{[HCN]}$$
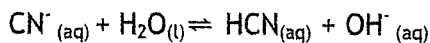
   b) $C_2H_3O_2^-$ (acetate ion)

   $$C_2H_3O_2^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HC_2H_3O_{2\,(aq)} + OH^-_{(aq)}$$
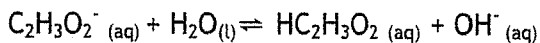
   $$K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$$

   c) $C_6H_5NH_2$ (aniline)

   $$C_6H_5NH_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons C_6H_5NH_3^+{}_{(aq)} + OH^-_{(aq)}$$
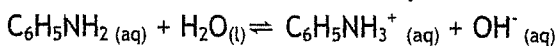
   $$K_b = \frac{[C_6H_5NH_3^+][OH^-]}{[C_6H_5NH_2]}$$

   d) $H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$

   $K_b = [H_3O^+][OH^-]$

4. Find the pH and % ionization of a 0.065 M solution of formic acid, (K_a=1.8 x 10^{-4}). [2.47, 5.2%]

$$HCHO_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons CHO_{2\,(aq)}^- + H_3O^+_{(aq)} \qquad K_a = 1.8 \times 10^{-4}$$

MR	1	1	1	1
I	0.065	---	0	0
C	-x	---	+x	+x
E	0.065-x	----	x	x

0.065/K_a < 1000
approx will not work

$$K_a = \frac{[CHO_2^-][H_3O^+]}{[HCHO_2]}$$

$$1.8 \times 10^{-4} = \frac{x^2}{(0.065-x)}$$

$$1.2 \times 10^{-5} - 1.8 \times 10^{-4}x = x^2$$

$$x^2 + 1.8 \times 10^{-4}x - 1.2 \times 10^{-5} = 0$$

a= 1, b=1.8x10^{-4}, c=-1.2x10^{-5}

~~x_1=-3.6x10^{-3}~~, x_2=3.4x10^{-3}

$[H_3O^+]$=3.4x10^{-3} M
pH = -log$[H_3O^+]$
pH=2.47

% ionization= $[H_3O^+]$/$[Acid]_{initial}$ x100%
  = (3.4x10^{-3})/0.065 x100%
  = 5.2%

5. Find the pH of a 0.325 M acetic acid solution, given the K_a is 1.8 x 10^{-5}. [2.62]

$$HC_2H_3O_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons C_2H_3O_{2\,(aq)}^- + H_3O^+_{(aq)} \qquad K_a=1.8 \times 10^{-5}$$

MR	1	1	1	1
I	0.325	---	0	0
C	-x	---	+x	+x
E	0.325-x	----	x	x

0.325/K_a > 1000
approx will work

$$K_a = \frac{[C_2H_3O_2^-][H_3O^+]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{x^2}{(0.325-x)}$$

$$5.9 \times 10^{-6} = x^2$$

$$x = 2.4 \times 10^{-3}$$

$[H_3O^+]$=2.4x10^{-3} M
pH = -log$[H_3O^+]$
pH=2.62

% ionization= $[H_3O^+]$/$[Acid]_{initial}$ x100%
  = (2.4x10^{-3})/0.325 x100%
  = 0.74 %

6. Find the pH of a solution that contains 0.0034 M lactic acid($HC_3H_5O_3$) (K_a =1.4 x 10^{-4}).

$$HC_3H_5O_{3\,(aq)} + H_2O_{(l)} \rightleftharpoons C_3H_5O_{3\,(aq)}^- + H_3O^+_{(aq)} \qquad K_a=1.4 \times 10^{-4}$$

MR	1	1	1	1
I	0.0034	---	0	0
C	-x	---	+x	+x
E	0.0034-x	----	x	x

0.0034/K_a < 1000
approx will not work

$$K_a = \frac{[C_3H_5O_3^-][H_3O^+]}{[HC_3H_5O_3]}$$

$$1.4 \times 10^{-4} = \frac{x^2}{(0.0034-x)}$$

$$x^2 + 1.4 \times 10^{-4}x - 4.8 \times 10^{-7} = 0$$

a= 1, b=1.4x10^{-4}, c= -4.8x10^{-7}

x_1=6.3x10^{-4}, x_2= ~~-7.7x10^{-4}~~

$[H_3O^+]$=6.3x10^{-4} M
pH = -log$[H_3O^+]$
pH=3.20

7. Find the hydronium ion concentration and pH for a 0.056 M solution of propanoic acid (K_a = 1.4 x 10^{-5}). [H_3O^+=8.9x10^{-4} pH=3.05]

$$HC_3H_5O_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons C_3H_5O_2^-{}_{(aq)} + H_3O^+{}_{(aq)} \qquad K_a = 1.4 \times 10^{-5}$$

MR	1	1	1	1
I	0.056	---	0	0
C	-x	---	+x	+x
E	0.056-x	----	x	x

$0.056/K_a > 1000$
approx will work

$$K_a = \frac{[C_3H_5O_2^-][H_3O^+]}{[HC_3H_5O_2]}$$

$$1.4 \times 10^{-5} = \frac{x^2}{(0.056 - \cancel{x})}$$

$$x = 8.9 \times 10^{-4}$$

$[H_3O^+]$=8.9x10^{-4} M
pH = -log[H_3O^+]
pH=3.05

8. Find the pH of a 0.600 M solution of methylamine CH_3NH_2. K_b = 4.4 x 10^{-4}. [12.20]

$$CH_3NH_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons CH_3NH_3^+{}_{(aq)} + OH^-{}_{(aq)} \qquad K_b = 4.4 \times 10^{-4}$$

MR	1	1	1	1
I	0.60	---	0	0
C	-x	---	+x	+x
E	0.60-x	----	x	x

$0.60/K_a > 1000$
approx will work

$$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$$

$$4.4 \times 10^{-4} = \frac{x^2}{(0.60 - \cancel{x})}$$

$$x = 1.6 \times 10^{-2}$$

$[OH^-]$=1.6x10^{-2} M
pOH = -log[OH^-]
pOH = 1.80
pH=12.20

9. What is the pH of a 5.6x10^{-4} M butanoic acid solution, given the pK_a = 4.82. [4.08]

$$HC_4H_7O_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons C_4H_7O_2^-{}_{(aq)} + H_3O^+{}_{(aq)}$$

MR	1	1	1	1
I	5.6x10^{-4}	---	0	0
C	-x	---	+x	+x
E	5.6x10^{-4} -x	----	x	x

5.6x$10^{-4}/K_a > 1000$
approx will work

pK_a=4.82
K_a=$10^{-4.82}$
K_a=1.5 x 10^{-5}

$$K_a = \frac{[C_4H_7O_2^-][H_3O^+]}{[HC_4H_7O_2]}$$

$$1.5 \times 10^{-5} = \frac{x^2}{(5.6 \times 10^{-4} - x)}$$

$$x^2 + 1.5 \times 10^{-5} x - 8.4 \times 10^{-9} = 0$$

a= 1, b=1.5x10^{-5}, c= -8.4x10^{-9}
x_1=8.4x10^{-5}, x_2= ~~-9.9x10^{-5}~~

$[H_3O^+]$=8.4x10^{-5} M
pH = -log[H_3O^+]
pH=4.08

10. Calculate the $[H_3O^+]$, the pH and the % ionization for a 0.50 mol/L HCN solution. [1.8×10^{-5}, 4.74, 3.6×10^{-3}%]

$$HCN_{(aq)} + H_2O_{(l)} \rightleftharpoons CN^-_{(aq)} + H_3O^+_{(aq)}$$

MR	1	1	1	1
I	0.50	---	0	0
C	-x	---	+x	+x
E	0.50-x	----	x	x

$0.50/K_a > 1000$

approx will work

$$K_a = \frac{[CN^-][H_3O^+]}{[HCN]}$$

$$6.2\times10^{-10} = \frac{x^2}{(0.50-\cancel{x})}$$

$$x = 1.8\times10^{-5}$$

$[H_3O^+]=1.8\times10^{-5}$ M
pH = $-\log[H_3O^+]$
pH=4.74

11. Calculate the $[OH^-]$, the pH and the % ionization for 0.25 mol/L ammonia solution. [2.1×10^{-3}, 11.32, 0.84%]

$$NH_{3\,(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+_{(aq)} + OH^-_{(aq)} \quad K_b=1.8 \times 10^{-5}$$

MR	1	1	1	1
I	0.25	---	0	0
C	-x	---	+x	+x
E	0.25-x	----	x	x

$0.25/K_a > 1000$

approx will work

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$1.8\times10^{-5} = \frac{x^2}{(0.25-\cancel{x})}$$

$$x = 2.1\times10^{-3}$$

$[OH^-]=2.1\times10^{-3}$ M
pOH = $-\log[OH^-]$
pOH = 2.68
pH=11.32

% ionization= $[OH^-]/[Base]_{initial}$ x100%
= $(2.1\times10^{-3})/0.25$ x100%
= **0.84 %**

12. Hydrazine, N_2H_4, has been used as a rocket fuel. Like ammonia, it is a weak base. A 0.15 M solution has a pH of 10.70. What is the K_b and pK_b for hydrazine and the pK_a of its conjugate acid? [$K_b=1.7\times10^{-6}$, pK_b =5.77, $pK_{a\,Conj}$ =8.23]

pH=10.70   pOH = 3.30 $[OH^-]=10^{-3.30}$ $[OH^-]$= 5.0×10^{-4}

$$N_2H_{4\,(aq)} + H_2O_{(l)} \rightleftharpoons N_2H_5^+_{(aq)} + OH^-_{(aq)} \quad K_b= ?$$

MR	1	1	1	1
I	0.15	---	0	0
C	-x	---	+x	+x
E	0.15-x =0.15	----	x =5.0×10^{-4}	x =5.0×10^{-4}

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

$$K_b = \frac{(5.0\times10^{-4})^2}{(0.15)}$$

$$K_b = 1.7\times10^{-6}$$

pK_b =$-\log K_b$
pK_b =5.77
$pK_{a\,Conj}$=14- pK_b
$pK_{a\,Conj}$ =8.23

13. If the pH of a weak base solution is 9.5 and the original concentration of base was 0.30 M what is the pOH, the concentration of OH⁻, the equilibrium concentration of the base and the K_b of the base? [pOH = 4.50, [OH⁻]= 3.2×10^{-5}, [B]=0.30 M, K_b=3.4×10^{-9}]

pH=9.50   pOH = 4.50   [OH⁻]=$10^{-4.50}$ [OH⁻]= 3.2×10^{-5}

	B $_{(aq)}$ +	H₂O$_{(l)}$ ⇌	BH⁺ $_{(aq)}$ +	OH⁻$_{(aq)}$
MR	1	1	1	1
I	0.30	---	0	0
C	-x	---	+x	+x
E	0.30-x =0.30	----	x =3.2x10⁻⁵	x =3.2x10⁻⁵

K_b= ?

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

$$K_b = \frac{(3.2 \times 10^{-5})^2}{(0.30)}$$

$$K_b = 3.4 \times 10^{-9}$$

14. If the pH of a hypochlorous acid solution is 4.17, what is the initial concentration? [0.16 M]

pH=4.17   [H₃O⁺]=$10^{-4.17}$ [H₃O⁺]=6.8×10^{-5}

	HClO $_{(aq)}$ +	H₂O$_{(l)}$ ⇌	ClO⁻ $_{(aq)}$ +	H₃O⁺$_{(aq)}$
MR	1	1	1	1
I	I	---	0	0
C	-x	---	+x	+x
E	I-x = I-6.8x10⁻⁵	----	x =6.8x10⁻⁵	x =6.8x10⁻⁵

K_a=2.9×10^{-8}

$$K_a = \frac{[ClO^-][H_3O^+]}{[HClO]}$$

$$2.9 \times 10^{-8} = \frac{(6.8 \times 10^{-5})^2}{(I - 6.8 \times 10^{-5})}$$

$$2.9 \times 10^{-8} I - 2.0 \times 10^{-12} = 4.6 \times 10^{-9}$$

$$I = 0.16$$

15. If the pH of a solution is equal to 8.70 and the K_b=9.6×10^{-7} what was the original concentration of the base? [3.1×10^{-5} M]

pH=8.70   pOH = 5.30   [OH⁻]=$10^{-5.30}$ [OH⁻]= 5.0×10^{-6}

	B $_{(aq)}$ +	H₂O$_{(l)}$ ⇌	BH⁺ $_{(aq)}$ +	OH⁻$_{(aq)}$
MR	1	1	1	1
I	I	---	0	0
C	-x	---	+x	+x
E	I-x =I-5.0x10⁻⁶	----	x =5.0x10⁻⁶	x =5.0x10⁻⁶

K_b= 9.6×10^{-7}

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

$$9.6 \times 10^{-7} = \frac{(5.0 \times 10^{-6})^2}{(I - 5.0 \times 10^{-6})}$$

$$9.6 \times 10^{-7} I - 4.8 \times 10^{-12} = 2.5 \times 10^{-11}$$

$$I = 3.1 \times 10^{-5}$$

16. A nitrous acid solution is 1.34% ionized at equilibrium. What is the hydronium ion concentration, pH and initial acid concentration? [$[H_3O^+]$=0.0531 M, pH=1.274, $[HNO_2]_{initial}$=3.96 M]

$$HNO_2 \text{ (aq)} + H_2O_{(l)} \rightleftharpoons NO_2^- \text{ (aq)} + H_3O^+ \text{ (aq)} \quad K_a=7.2\times10^{-4}$$

MR	1	1	1	1
I	I	---	0	0
C	-x	---	+x	+x
E	I-x =0.9866I	----	x =0.0134I	x =0.0134I

Since the acid is 1.34% ionized at equilibrium,
$$x=0.0134I$$

$$K_a = \frac{[NO_2^-][H_3O^+]}{[HNO_2]}$$

$$7.2\times10^{-4} = \frac{(0.0134I)^2}{(0.9866I)}$$

$$7.2\times10^{-4} = \frac{1.80\times10^{-4}I^2}{0.9866I}$$

$$7.2\times10^{-4} = 1.82\times10^{-4}I$$

$$I = 3.96M$$

x=0.0134I
x=(0.0134)(3.96)
x=0.0531 M
$[H_3O^+]$=0.0531 M
pH = -log$[H_3O^+]$
pH=1.274

17. A solution of ethylamine is 1.25% ionized at equilibrium. What is the hydroxide ion concentration and initial base concentration? [$[OH^-]$=0.0340 M, $[CH_3CH_2NH_2]_{initial}$=2.72 M]

$$CH_3CH_2NH_2 \text{ (aq)} + H_2O_{(l)} \rightleftharpoons CH_3CH_2NH_3^+ \text{ (aq)} + OH^- \text{ (aq)} \quad K_b=4.3\times10^{-4}$$

MR	1	1	1	1
I	I	---	0	0
C	-x	---	+x	+x
E	I-x =0.9875I	----	x =0.0125I	x =0.0125I

Since the base is 1.25% ionized at equilibrium,
$$x=0.0125I$$

$$K_b = \frac{[CH_3CH_2NH_3^+][OH^-]}{[CH_3CH_2NH_2]}$$

$$4.3\times10^{-4} = \frac{(0.0125I)^2}{0.9875I}$$

$$4.3\times10^{-4} = \frac{1.56\times10^{-4}I^2}{0.9875I}$$

$$4.3\times10^{-4} = 1.58\times10^{-4}I$$

$$I = 2.72M$$

x=0.0125I
x=(0.0125)(2.72)
x=0.0340 M
$[OH^-]$=0.0340 M

18. Calculate the pH of the following solutions,

a) 0.36 M sodium acetate solution [9.15]

$$C_2H_3O_2^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HC_2H_3O_2{}_{(aq)} + OH^-{}_{(aq)}$$

	$C_2H_3O_2^-$	H_2O	$HC_2H_3O_2$	OH^-
MR	1	1	1	1
I	0.36	---	0	0
C	-x	---	+x	+x
E	0.36-x	----	X	x

$K_b = K_w/K_{a\ Acid}$

0.36/K_b > 1000
approx will work

$$K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$$

$$\frac{K_W}{K_{aAcid}} = \frac{x^2}{0.36 - \cancel{x}}$$

$$5.6 \times 10^{-10} = \frac{x^2}{0.36}$$

$$x = 1.4 \times 10^{-5}$$

$[OH^-]=1.4\times10^{-5}$ M
pOH = -log[OH⁻]
pOH = 4.85
pH=9.15

b) 0.25 M ammonium nitrate [4.92]

$$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_3{}_{(aq)} + H_3O^+{}_{(aq)}$$

	NH_4^+	H_2O	NH_3	H_3O^+
MR	1	1	1	1
I	0.25	---	0	0
C	-x	---	+x	+x
E	0.25-x	----	X	X

$K_a = K_w/K_{b\ Base}$

0.25/K_a > 1000
approx will work

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_3]}$$

$$\frac{K_W}{K_{bBase}} = \frac{x^2}{0.25 - \cancel{x}}$$

$$5.6 \times 10^{-10} = \frac{x^2}{0.25}$$

$$x = 1.2 \times 10^{-5}$$

$[H_3O^+]=1.2\times10^{-5}$ M
pH = -log[H₃O⁺]
pH=4.92

c) 0.16 M calcium formate (Ca(CH₂O)₂) solution [8.48]

$$CHO_2^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HCHO_2{}_{(aq)} + OH^-{}_{(aq)}$$

	CHO_2^-	H_2O	$HCHO_2$	OH^-
MR	1	1	1	1
I	0.32	---	0	0
C	-x	---	+x	+x
E	0.32-x	----	x	X

$K_b = K_w/K_{a\ Acid}$

0.16/K_b > 1000
approx will work

$$K_b = \frac{[HCH_2O_2][OH^-]}{[CH_2O_2^-]}$$

$$\frac{K_W}{K_{aAcid}} = \frac{x^2}{0.32 - \cancel{x}}$$

$$5.6 \times 10^{-11} = \frac{x^2}{0.32}$$

$$x = 4.2 \times 10^{-6}$$

$[OH^-]=4.2\times10^{-6}$ M
pOH = -log[OH⁻]
pOH = 5.38
pH= 8.62

d) 0.45 M potassium hypochlorite   [10.59]

$$ClO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HClO_{(aq)} + OH^-_{(aq)} \quad K_b = K_w/K_{a\,Acid}$$

MR	1	1	1	1
I	0.45	---	0	0
C	-x	---	+x	+x
E	0.45-x	----	x	x

$0.45/K_b > 1000$
approx will work

$$K_b = \frac{[HClO][OH^-]}{[ClO^-]}$$

$$\frac{K_W}{K_{aAcid}} = \frac{x^2}{0.45 - \cancel{x}}$$

$$3.4 \times 10^{-7} = \frac{x^2}{0.45}$$

$$x = 3.9 \times 10^{-4}$$

$[OH^-]=3.9\times10^{-4}$ M
pOH = -log$[OH^-]$
pOH = 3.41
pH= 10.59

19. A 27.0 mL sample of 0.28 M acetic acid is titrated with 0.15 M calcium hydroxide.
    a) Calculate the volume of calcium hydroxide required to reach the equivalence point. [25mL]
    b) Calculate the pH at equivalence.   [8.96]

$$2\,HC_2H_3O_{2\,(aq)} + Ca(OH)_{2\,(aq)} \rightarrow Ca(C_2H_3O_2)_{2\,(aq)} + 2\,H_2O_{(l)}$$

	$2\,HC_2H_3O_2$	$Ca(OH)_2$	$Ca(C_2H_3O_2)_2$
c	0.28	0.15	0.073 ($[C_2H_3O_2^-]$=0.15 M)
V	0.027	0.025	0.052
n	7.6×10^{-3}	3.8×10^{-3}	3.8×10^{-3}

↳need 7.6×10^{-3} mol OH $\times \frac{1\,mol\ Ca(OH)_2}{2\,mol\ OH} = 3.8\times10^{-3}$ mol $Ca(OH)_2$

$$C_2H_3O_2^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HC_2H_3O_{2(aq)} + OH^-_{(aq)}$$

MR	1	1	1	1
I	0.15	--	0	0
C	-x	---	x	x
E	0.15-x	--	x	x

$0.15/K_b > 1000$
approx will work

$$K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$$

$$\frac{K_W}{K_{aAcid}} = \frac{x^2}{0.15 - \cancel{x}}$$

$$5.6 \times 10^{-10} = \frac{x^2}{0.15}$$

$$x = 9.2 \times 10^{-6}$$

$[OH^-]=9.2\times10^{-6}$
pOH = -log$[OH^-]$
pOH=5.04
pH=8.96

20. A 175 mL sample of 0.30 M ammonia is titrated with 0.45 M hydrochloric acid.
    a) Calculate the volume of hydrochloric acid required to reach the equivalence point. [120mL]
    b) Calculate the pH at equivalence. [9.00]

	HCl $_{(aq)}$ +	NH$_{3 (aq)}$ →	NH$_4$Cl $_{(aq)}$
c	0.45	0.30	0.18 ([NH$_4^+$]=0.18 M)
V	0.12	0.175	0.295
n	0.053	0.053	0.053

NH$_4^+$ $_{(aq)}$ + H$_2$O$_{(l)}$ ⇌ NH$_{3(aq)}$ + H$_3$O$^+$ $_{(aq)}$

MR	1	1	1	1
I	0.18	--	0	0
C	-x	---	X	X
E	0.18-x	--	X	X

0.18/K$_a$ > 1000
approx will work

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_3]}$$

$$\frac{K_W}{K_{bBase}} = \frac{x^2}{0.18 - x}$$

$$5.6 \times 10^{-10} = \frac{x^2}{0.18}$$

$$x = 1.0 \times 10^{-5}$$

[H$_3$O$^+$]=1.0x10^{-5}
pH=-log[H$_3$O$^+$]
pH=5.00

21. A 225 mL sample of 0.24 M formic acid is titrated with 0.30 M sodium hydroxide.
    a) Calculate the volume of hydrochloric acid required to reach the equivalence point. [180mL]
    b) Calculate the pH at equivalence. [8.43]

	HCHO$_2$ $_{(aq)}$ +	NaOH $_{(aq)}$ →	NaCHO$_2$ $_{(aq)}$ + H$_2$O $_{(l)}$
c	0.24	0.30	0.13 ([CHO$_2^-$]=0.13 M]
V	0.225	0.180	0.405 CHO$_2$
n	0.054	0.054	0.054

CHO$_2^-$ $_{(aq)}$ + H$_2$O$_{(l)}$ ⇌ HCHO$_{2(aq)}$ + OH$^-$ $_{(aq)}$

MR	1	1	1	1
I	0.13	--	0	0
C	-x	---	X	X
E	0.13-x	--	X	X

0.13/K$_b$ > 1000
approx will work

$$K_b = \frac{[HCHO_2][OH^-]}{[CHO_2^-]}$$

$$\frac{K_W}{K_{aAcid}} = \frac{x^2}{0.13 - x}$$

$$5.6 \times 10^{-11} = \frac{x^2}{0.13}$$

$$x = 2.7 \times 10^{-6}$$

[OH$^-$]=2.7x10^{-6}
pOH = -log[OH$^-$]
pOH =5.57
pH=8.43

22. A 180 mL sample of 0.16 M hydrazine is titrated with 0.12 M nitric acid.
    a) Calculate the volume of nitric acid required to reach the equivalence point.   [240mL]
    b) Calculate the pH at equivalence.   [6.08]

	$HNO_{3 (aq)}$ +	$N_2H_{4 (aq)}$ $\rightarrow$	$N_2H_5NO_{3 (aq)}$
c	0.12	0.16	0.069 ($[N_2H_5^+]$=0.069 M)
V	0.24	0.180	0.42
n	0.029	0.029	0.029

	$N_2H_5^+{}_{(aq)}$ +	$H_2O_{(l)}$ $\rightleftharpoons$	$N_2H_{4 (aq)}$ +	$H_3O^+{}_{(aq)}$
MR	1	1	1	1
I	0.069	--	0	0
C	-x	---	x	x
E	0.069-x	--	x	x

0.069/K_a > 1000
approx will work

$$K_a = \frac{[N_2H_4][H_3O^+]}{[N_2H_5^+]}$$

$$\frac{K_W}{K_{bBase}} = \frac{x^2}{0.069 - x}$$

$$1.0 \times 10^{-11} = \frac{x^2}{0.069}$$

$$x = 8.3 \times 10^{-7}$$

$[H_3O^+]$=8.3x10^{-5}
pH=6.08

23. A 375 mL sample of 0.25 M hydrofluoric acid is titrated with 0.15 M potassium hydroxide.
    a) the volume of potassium hydroxide needed to reach the equivalence point   [630 mL]
    b) the pH at equivalence.   [8.07]

	$HF_{(aq)}$ +	$KOH_{(aq)}$ $\rightarrow$	$KF_{(aq)}$ +	$H_2O_{(l)}$
c	0.25	0.15	0.094($[F^-]$=0.094 M]	
V	0.375	0.63	1.005	
n	0.094	0.094	0.094	

	$F^-{}_{(aq)}$ +	$H_2O_{(l)}$ $\rightleftharpoons$	$HF_{(aq)}$ +	$OH^-{}_{(aq)}$
MR	1	1	1	1
I	0.094	--	0	0
C	-x	---	x	x
E	0.094-x	--	x	x

$K_{a\ HF}$=6.6x10^{-4}

0.094/K_b > 1000
approx will work

$$K_b = \frac{[HF][OH^-]}{[F^-]}$$

$$\frac{K_W}{K_{aAcid}} = \frac{x^2}{0.094 - x}$$

$$1.5 \times 10^{-11} = \frac{x^2}{0.094}$$

$$x = 1.2 \times 10^{-6}$$

$[OH^-]$=1.2x10^{-6}
pOH = -log$[OH^-]$
pOH =5.92
pH=8.08

24. If 125 mL of 0.15 M hydrochloric acid is needed for 95 mL of ammonia to reach equivalence, find:
    a) the concentration of the original ammonia solution.   [0.20 M]
    b) the pH of the solution at the equivalence point.   [5.16]

$$HCl_{(aq)} + \quad NH_{3\,(aq)} \rightarrow \quad NH_4Cl_{(aq)}$$

	HCl	NH₃	NH₄Cl
c	0.15	0.20	0.086 ([NH₄⁺]=0.086 M)
V	0.125	0.095	0.22
n	0.019	0.019	0.019

$$NH_4^+{}_{(aq)} + \quad H_2O_{(l)} \rightleftharpoons \quad NH_{3(aq)} + H_3O^+{}_{(aq)}$$

$K_{b\ ammonia}=1.8\times10^{-5}$

	NH_4^+	H_2O	NH_3	H_3O^+
MR	1	1	1	1
I	0.086	--	0	0
C	-x	---	x	x
E	0.086-x	--	x	x

$0.086/K_a > 1000$
approx will work

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_3]}$$

$$\frac{K_W}{K_{bBase}} = \frac{x^2}{0.086 - \cancel{x}}$$

$$5.6\times10^{-10} = \frac{x^2}{0.086}$$

x=6.9x10⁻⁶
[H₃O⁺]=6.9x10⁻⁶
pH=-log[H₃O⁺]
pH=5.16

25. If 230 mL of 0.35 M cyanic acid (K_a=3.5x10⁻⁴) is titrated with 0.25 M calcium hydroxide. Calculate:
    a) the pH of the acid sample before titration.   [1.96]
    b) the volume of calcium hydroxide needed to reach the equivalence point. [160 mL]
    c) the pH of the solution at the equivalence point.   [8.39]

a) $$HCNO_{(aq)} + H_2O_{(l)} \rightleftharpoons CNO^-{}_{(aq)} + H_3O^+{}_{(aq)}$$

$K_a = 3.5\times10^{-4}$

	$HCNO$	H_2O	CNO^-	H_3O^+
MR	1	1	1	1
I	0.35	---	0	0
C	-x	---	+x	+x
E	0.35-x	----	x	x

$0.35/K_a = 1000$
approx will work

$$K_a = \frac{[CNO^-][H_3O^+]}{[HCNO]}$$

$$3.5\times10^{-4} = \frac{x^2}{(0.35-\cancel{x})}$$

x=1.1x10⁻²
[H₃O⁺]=1.1x10⁻²
pH=1.96

b) and c) $$2\ HCNO_{(aq)} + Ca(OH)_{2\,(aq)} \rightarrow Ca(CNO)_{2\,(aq)} + 2\ H_2O_{(l)}$$

	HCNO	Ca(OH)₂	Ca(CNO)₂
c	0.35	0.25	0.10 ([CNO⁻]=0.20 M)
V	0.23	0.16	0.39
n	0.081	0.0405	0.0405

$$CNO^-{}_{(aq)} + \quad H_2O_{(l)} \rightleftharpoons \quad HCNO_{(aq)} + OH^-{}_{(aq)}$$

	CNO^-	H_2O	$HCNO$	OH^-
MR	1	1	1	1
I	0.20	--	0	0
C	-x	---	x	x
E	0.20-x	--	x	x

$0.20/K_b > 1000$
approx will work

$$K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$$

$$\frac{K_W}{K_{aAcid}} = \frac{x^2}{0.20 - \cancel{x}}$$

$$2.9\times10^{-11} = \frac{x^2}{0.20}$$

x=2.4x10⁻⁶
[OH⁻]=2.4x10⁻⁶
pOH = -log[OH⁻]
pOH=5.62
pH=8.38

26. A 65 mL sample of 0.45 M methylamine is titrated with 0.45 M nitric acid. Calculate:
  a) the pH of the base sample before titration. [12.15]
  b) the volume of nitric acid needed to reach the equivalence point. [65 mL]
  c) the pH of the solution at the equivalence point. [5.66]

a) pH before titration

$CH_3NH_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons CH_3NH_3^+{}_{(aq)} + OH^-{}_{(aq)}$ $K_b = 4.4 \times 10^{-4}$

MR	1	1	1	1
I	0.45	---	0	0
C	-x	---	+x	+x
E	0.45-x	----	x	x

$0.45/K_b > 1000$
approx will work

$$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$$

$$4.4 \times 10^{-4} = \frac{x^2}{(0.45 - \cancel{x})}$$

$x = 1.4 \times 10^{-2}$
$[OH^-] = 1.4 \times 10^{-2}$ M
pOH = -log[OH^-]
pOH = 1.85
pH = 12.15

b) and c) $HNO_{3\,(aq)} + CH_3NH_{2\,(aq)} \rightarrow CH_3NH_3NO_{3\,(aq)} + 2\,H_2O_{(l)}$

c	0.45	0.45	0.22 ($[CH_3NH_3^+] = 0.22$ M)
V	0.65	0.65	0.13
n	0.029	0.029	0.029

$CH_3NH_3^+ + H_2O_{(l)} \rightleftharpoons CH_3NH_{2\,(aq)} + H_3O^+{}_{(aq)}$

MR	1	1	1	1
I	0.22	--	0	0
C	-x	---	x	x
E	0.22-x	--	x	x

$0.22/K_a > 1000$
approx will work

$$K_a = \frac{[CH_3NH_2][H_3O^+]}{[CH_3NH_3^+]}$$

$$\frac{K_W}{K_{bBase}} = \frac{x^2}{0.22 - \cancel{x}}$$

$$2.3 \times 10^{-11} = \frac{x^2}{0.22}$$

$$x = 2.2 \times 10^{-6}$$

$[H_3O^+] = 2.2 \times 10^{-6}$
pH = -log[H_3O^+]
pH = 5.66

27. If 65 mL of 0.20 M calcium hydroxide is need to titrate 85 mL of an unknown concentration of nitrous acid to equivalence point, calculate:
    a) the concentration of the original nitrous acid solution.   [0.31 M]
    b) the pH of the solution at the equivalence point.   [8.19]

$2\ HNO_{2\ (aq)}\ +\ Ca(OH)_{2\ (aq)} \rightarrow Ca(NO_2)_{2\ (aq)}\ +\ 2\ H_2O_{\ (l)}$

c	0.31	0.20	0.087 ([NO$_2^-$]=0.17 M)
V	0.085	0.065	0.15
n	0.026	0.013	0.013

$NO_{2\ aq)}^-\ +\ H_2O_{(l)} \rightleftharpoons HNO_{2\ (aq)}\ +\ OH^-_{\ (aq)}$ $K_a = 7.2\times10^{-4}$

MR	1	1	1	1
I	0.17	--	0	0
C	-x	---	x	x
E	0.17-x	--	x	x

0.17/K$_b$ > 1000
approx will work

$$K_b = \frac{[HNO_2][OH^-]}{[NO_2^-]}$$

$$\frac{K_W}{K_{aAcid}} = \frac{x^2}{0.17 - x}$$

$$1.4\times10^{-11} = \frac{x^2}{0.17}$$

x=1.5x10^{-6}
[OH$^-$]=1.5x10^{-6}
pOH = -log[OH$^-$]
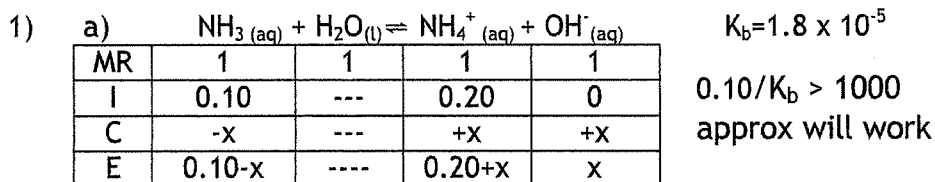pOH=5.81
pH=8.19

Buffer Questions

For each of the following questions calculate the pH of a) the initial buffer, b) after the addition of the H$_3$O$^+$ to the initial buffer and c) after the addition of the OH$^-$ to the initial buffer.

Question	Acid or Base	Conjugate	Volume (L)	moles of H$_3$O$^+$	moles of OH$^-$
1	0.300 mol of NH$_3$	0.600 mol of NH$_4$NO$_3$	3.0	0.150	0.150
2	0.400 mol HC$_2$H$_3$O$_2$	0.400 mol NaC$_2$H$_3$O$_2$	2.0	0.200	0.100
3	0.250 mol H$_2$CO$_3$	0.200 mol NaHCO$_3$	1.0	0.050	0.060
4	0.200 mol CH$_3$NH$_2$	0.150 mol CH$_3$NH$_3$NO$_3$	3.0	0.020	0.030
5	0.900 mol HCN	0.800 mol NaCN	2.0	0.200	0.300
6	1.50 mol benzoic acid	1.20 mol sodium benzoate	3.0	0.600	0.300
7	0.450 mol HClO	0.350 mol KClO	1.0	0.250	0.15

Answers

Question	Initial pH	After H$_3$O$^+$	After OH$^-$
1	8.95	8.56	9.26
2	4.74	4.27	4.96
3	6.26	6.06	6.49
4	10.76	10.65	10.92
5	9.15	8.96	9.47
6	4.10	3.66	4.30
7	7.43	6.70	7.77

1) a) $NH_{3\ (aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$ $K_b = 1.8 \times 10^{-5}$

MR	1	1	1	1
I	0.10	---	0.20	0
C	-x	---	+x	+x
E	0.10-x	----	0.20+x	x

$0.10/K_b > 1000$

approx will work

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

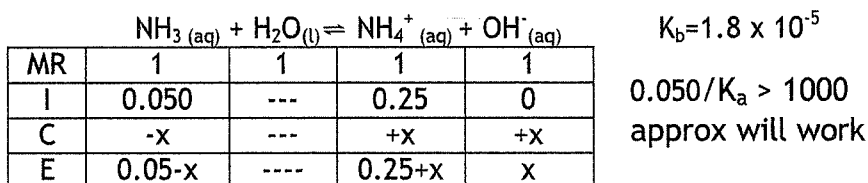$$1.8 \times 10^{-5} = \frac{(0.20 + x)x}{(0.10 - x)}$$

$$x = 9.0 \times 10^{-6}$$

[OH⁻]=9.0x10⁻⁶ M
pOH = -log[OH⁻]
pOH = 5.05
pH=8.95

b) If you add 0.150 mol of H_3O^+ in 3.0 L, the equilibrium will shift right, therefore [NH₃] will decrease to 0.050 M and [NH₄⁺] will increase to 0.25 M

$NH_{3\ (aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$ $K_b = 1.8 \times 10^{-5}$

MR	1	1	1	1
I	0.050	---	0.25	0
C	-x	---	+x	+x
E	0.05-x	----	0.25+x	x

$0.050/K_a > 1000$

approx will work

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

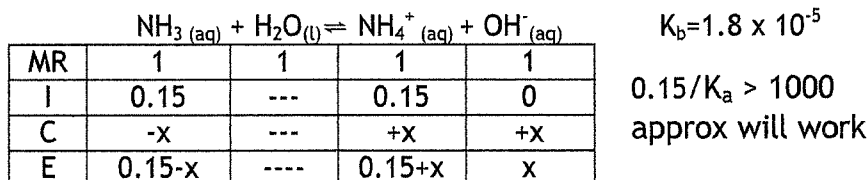$$1.8 \times 10^{-5} = \frac{(0.25 + x)x}{(0.05 - x)}$$

$$x = 3.6 \times 10^{-6}$$

[OH⁻]=3.6x10⁻⁶ M
pOH = -log[OH⁻]
pOH = 5.44
pH=8.56

c) If you add 0.150 mol of OH⁻ in 3.0 L, the equilibrium will shift left, therefore [NH₃] will increase to 0.150 M and [NH₄⁺] will decrease to 0.15 M

$NH_{3\ (aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$ $K_b = 1.8 \times 10^{-5}$

MR	1	1	1	1
I	0.15	---	0.15	0
C	-x	---	+x	+x
E	0.15-x	----	0.15+x	x

$0.15/K_a > 1000$

approx will work

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$1.8 \times 10^{-5} = \frac{(0.15 + x)x}{(0.15 - x)}$$

$$x = 1.8 \times 10^{-5}$$

[OH⁻]=3.6x10⁻⁶ M
pOH = -log[OH⁻]
pOH = 4.74
pH=9.26

2) a) $HC_2H_3O_{2\ (aq)} + H_2O_{(l)} \rightleftharpoons C_2H_3O_2^-{}_{(aq)} + H_3O^+{}_{(aq)}$ $K_a=1.8 \times 10^{-5}$
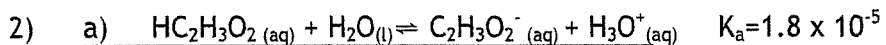
MR	1	1	1	1
I	0.20	---	0.20	0
C	-x	---	+x	+x
E	0.20-x	----	0.20+x	x

$0.20/K_a > 1000$
approx will work

$$K_a = \frac{[C_2H_3O_2^-][H_3O^+]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{(0.20 + \cancel{x})x}{(0.20 - \cancel{x})}$$

$$x = 1.8 \times 10^{-5}$$

$[H_3O^+]=1.8\times10^{-5}$
$pH=-\log[H_3O^+]$
pH=4.74

b) If you add 0.200 mol H_3O^+ in 2.0 L, the equilibrium will shift left, therefore $[HC_2H_3O_2]$ will increase to 0.30 M and $[C_2H_3O_2^-]$ will decrease to 0.10 M

$HC_2H_3O_{2\ (aq)} + H_2O_{(l)} \rightleftharpoons C_2H_3O_2^-{}_{(aq)} + H_3O^+{}_{(aq)}$ $K_a=1.8 \times 10^{-5}$
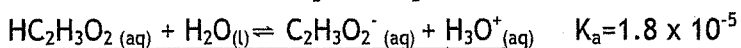
MR	1	1	1	1
I	0.30	---	0.10	0
C	-x	---	+x	+x
E	0.30-x	----	0.10+x	x

$0.10/K_a > 1000$
approx will work

$$K_a = \frac{[C_2H_3O_2^-][H_3O^+]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{(0.10 + \cancel{x})x}{(0.30 - \cancel{x})}$$

$$x = 5.4 \times 10^{-5}$$

$[H_3O^+]=5.4\times10^{-5}$
$pH=-\log[H_3O^+]$
pH=4.27

c) If you add 0.100 mol of OH^- in 2.0 L, the equilibrium will shift right, therefore $[HC_2H_3O_2]$ will decrease to 0.150 M and $[C_2H_3O_2^-]$ will increase to 0.25 M

$HC_2H_3O_{2\ (aq)} + H_2O_{(l)} \rightleftharpoons C_2H_3O_2^-{}_{(aq)} + H_3O^+{}_{(aq)}$ $K_a=1.8 \times 10^{-5}$
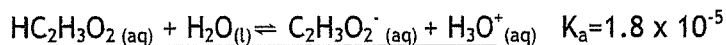
MR	1	1	1	1
I	0.15	---	0.25	0
C	-x	---	+x	+x
E	0.15-x	----	0.25+x	x

$0.15/K_a > 1000$
approx will work

$$K_a = \frac{[C_2H_3O_2^-][H_3O^+]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{(0.25 + \cancel{x})x}{(0.15 - \cancel{x})}$$

$$x = 1.1 \times 10^{-5}$$

$[H_3O^+]=1.1\times10^{-5}$
$pH=-\log[H_3O^+]$
pH=4.96

3) a)

$H_2CO_{3\,(aq)}$ +	$H_2O_{(l)}$	$\rightleftharpoons$ $HCO_{3\,(aq)}^-$ +	$H_3O^+_{(aq)}$	$K_a = 4.4 \times 10^{-7}$

MR	1	1	1	1
I	0.25	---	0.20	0
C	-x	---	+x	+x
E	0.25-x	----	0.20+x	x

$0.20/K_a > 1000$

approx will work
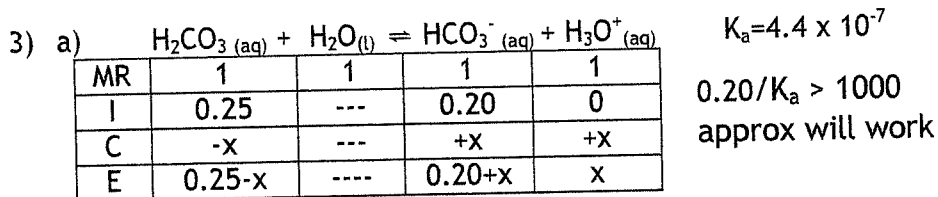
$$K_a = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]}$$

$$4.4 \times 10^{-7} = \frac{(0.20 + \cancel{x})x}{(0.25 - \cancel{x})}$$

$$x = 5.5 \times 10^{-7}$$

$[H_3O^+] = 5.5 \times 10^{-7}$
pH = -log$[H_3O^+]$
pH = 6.26

b) If you add 0. 050 mol H_3O^+ in 1.0 L, the equilibrium will shift left, therefore $[H_2CO_3]$ will increase to 0.30 M and $[HCO_3^-]$ will decrease to 0.15 M

$H_2CO_{3\,(aq)}$ +	$H_2O_{(l)}$	$\rightleftharpoons$ $HCO_{3\,(aq)}^-$ +	$H_3O^+_{(aq)}$	$K_a = 4.4 \times 10^{-7}$

MR	1	1	1	1
I	0.30	---	0.15	0
C	-x	---	+x	+x
E	0.30-x	----	0.15+x	x

$0.15/K_a > 1000$

approx will work
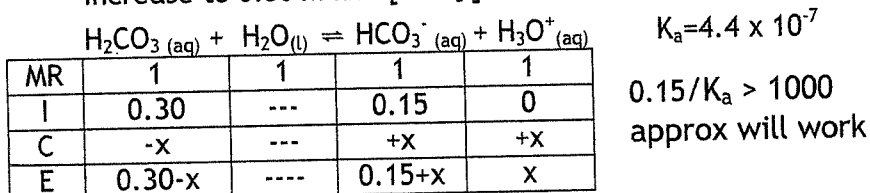
$$K_a = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]}$$

$$4.4 \times 10^{-7} = \frac{(0.15 + \cancel{x})x}{(0.30 - \cancel{x})}$$

$$x = 8.8 \times 10^{-7}$$

$[H_3O^+] = 8.8 \times 10^{-7}$
pH = -log$[H_3O^+]$
pH = 6.06

c) If you add 0.060 mol of OH^- in 1.0 L, the equilibrium will shift right, therefore $[H_2CO_3]$ will decrease to 0.150 M and $[HCO_3^-]$ will increase to 0.25 M

$H_2CO_{3\,(aq)}$ +	$H_2O_{(l)}$	$\rightleftharpoons$ $HCO_{3\,(aq)}^-$ +	$H_3O^+_{(aq)}$	$K_a = 4.4 \times 10^{-7}$

MR	1	1	1	1
I	0.19	---	0.26	0
C	-x	---	+x	+x
E	0.19-x	----	0.26+x	x

$0.19/K_a > 1000$
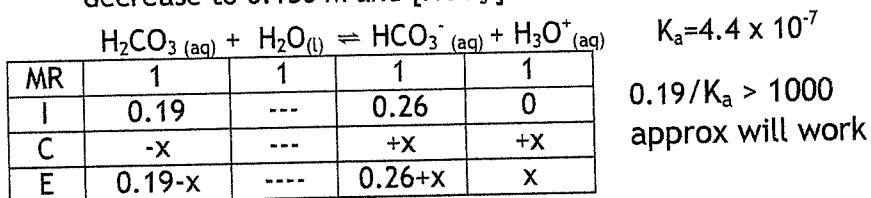
approx will work

$$K_a = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]}$$

$$4.4 \times 10^{-7} = \frac{(0.26 + \cancel{x})x}{(0.19 - \cancel{x})}$$

$$x = 3.2 \times 10^{-7}$$

$[H_3O^+] = 3.2 \times 10^{-7}$
pH = -log$[H_3O^+]$
pH = 6.49

4) a) $CH_3NH_2 {}_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3NH_3^+ {}_{(aq)} + OH^-{}_{(aq)}$ $K_b = 4.4 \times 10^{-4}$
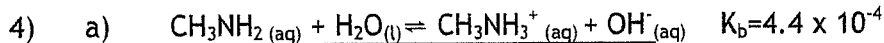
MR	1	1	1	1
I	0.067	---	0.05	0
C	-x	---	+x	+x
E	0.067-x	----	0.05+x	x

$0.05/K_b < 1000$

approx will not work

$$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$$

$$4.4 \times 10^{-4} = \frac{(0.050+x)x}{(0.067-x)}$$

$$x^2 + 0.050x - 2.9 \times 10^{-5} = 0$$

a=1, b=0.050, c= -2.9x10^{-5}

$x_1 = 5.7 \times 10^{-4}$ $\cancel{x_2 = -0.051}$

$[OH^-] = 5.7 \times 10^{-4}$ M

pOH = -log[OH$^-$]

pOH = 3.24

pH=10.76

b) If you add 0.020 mol of H_3O^+ in 3.0 L, the equilibrium will shift right, therefore $[CH_3NH_2]$ will decrease to 0.060 M and $[CH_3NH_3^+]$ will increase to 0.25 M

$CH_3NH_2 {}_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3NH_3^+ {}_{(aq)} + OH^-{}_{(aq)}$ $K_b = 4.4 \times 10^{-4}$
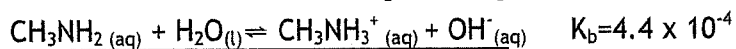
MR	1	1	1	1
I	0.060	---	0.057	0
C	-x	---	+x	+x
E	0.060-x	----	0.057+x	x

$0.057/K_b < 1000$

approx will not work

$$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$$

$$4.4 \times 10^{-4} = \frac{(0.057+x)x}{(0.060-x)}$$

$$x^2 + 0.057x - 2.6 \times 10^{-5} = 0$$

a=1, b=0.057, c= -2.6x10^{-5}

$x_1 = 4.5 \times 10^{-4}$ $\cancel{x_2 = -0.057}$

$[OH^-] = 4.5 \times 10^{-4}$ M

pOH = -log[OH$^-$]

pOH = 3.35

pH=10.65

c) If you add 0.030 mol of OH$^-$ in 3.0 L, the equilibrium will shift left, therefore $[CH_3NH_2]$ will increase to 0.077 M and $[CH_3NH_3^+]$ will decrease to 0.040 M

$CH_3NH_2 {}_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3NH_3^+ {}_{(aq)} + OH^-{}_{(aq)}$ $K_b = 4.4 \times 10^{-4}$
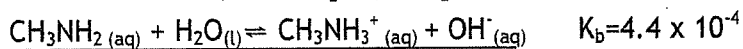
MR	1	1	1	1
I	0.077	---	0.040	0
C	-x	---	+x	+x
E	0.077-x	----	0.040+x	x

$0.040/K_b < 1000$

approx will not work

$$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$$

$$4.4 \times 10^{-4} = \frac{(0.040+x)x}{(0.077-x)}$$

$$x^2 + 0.040x - 3.4 \times 10^{-5} = 0$$

a=1, b=0.040, c= -3.4x10^{-5}

$x_1 = 8.3 \times 10^{-4}$ $\cancel{x_2 = -0.041}$

$[OH^-] = 8.3 \times 10^{-4}$ M

pOH = -log[OH$^-$]

pOH = 3.08

pH=10.92

a)

HCN $_{(aq)}$ + H$_2$O$_{(l)}$ ⇌		CN$^-$ $_{(aq)}$ +	H$_3$O$^+$ $_{(aq)}$
1	1	1	1
MR	---	0.40	0
I 0.45	---	+x	+x
C -x	----	0.40+x	x
E 0.45-x		0.40+x	x

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

$0.40/K_a > 1000$

approx will work

$$K_a = \frac{[CN^-][H_3O^+]}{[HCN]}$$

$$6.2 \times 10^{-10} = \frac{(0.40 + \cancel{x})x}{(0.45 - \cancel{x})}$$

$$x = 7.0 \times 10^{-10}$$

$[H_3O^+] = 7.0 \times 10^{-10}$
pH = -log$[H_3O^+]$
pH = 9.15

b) If you add 0.200 mol H$_3$O$^+$ in 2.0 L, the equilibrium will shift left, therefore [HCN] will increase to 0.55 M and [CN$^-$] will decrease to 0.30 M

HCN $_{(aq)}$ + H$_2$O$_{(l)}$ ⇌		CN$^-$ $_{(aq)}$ +	H$_3$O$^+$ $_{(aq)}$
1	1	1	1
MR	---	0.30	0
I 0.55	---	+x	+x
C -x	----	0.30+x	x
E 0.55-x		0.30+x	x

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

$0.30/K_a > 1000$

approx will work

$$K_a = \frac{[C_2H_3O_2^-][H_3O^+]}{[HC_2H_3O_2]}$$

$$6.2 \times 10^{-10} = \frac{(0.30 + \cancel{x})x}{(0.55 - \cancel{x})}$$

$$x = 1.1 \times 10^{-9}$$

$[H_3O^+] = 1.1 \times 10^{-9}$
pH = -log$[H_3O^+]$
pH = 8.96

c) If you add 0.300 mol of OH$^-$ in 2.0 L, the equilibrium will shift right, therefore [HCN] will decrease to 0.30 M and [CN$^-$] will increase to 0.55 M

HCN $_{(aq)}$ + H$_2$O$_{(l)}$ ⇌		CN$^-$ $_{(aq)}$ +	H$_3$O$^+$ $_{(aq)}$
1	1	1	1
MR	---	0.55	0
I 0.30	---	+x	+x
C -x	----	0.55+x	x
E 0.30-x		0.55+x	x

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

$0.30/K_a > 1000$

approx will work

$$K_a = \frac{[C_2H_3O_2^-][H_3O^+]}{[HC_2H_3O_2]}$$

$$6.2 \times 10^{-10} = \frac{(0.55 + \cancel{x})x}{(0.30 - \cancel{x})}$$

$$x = 3.4 \times 10^{-10}$$

$[H_3O^+] = 1.1 \times 10^{-5}$
pH = -log$[H_3O^+]$
pH = 9.47

6) a) $HC_7H_5O_{2\ (aq)} + H_2O_{(l)} \rightleftharpoons C_7H_5O_2^-{}_{(aq)} + H_3O^+{}_{(aq)}$ $K_a = 6.3 \times 10^{-5}$
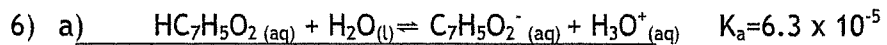
MR	1	1	1	1
I	0.50	---	0.40	0
C	-x	---	+x	+x
E	0.50-x	----	0.40+x	x

$0.40/K_a > 1000$

approx will work

$$K_a = \frac{[C_7H_5O_2^-][H_3O^+]}{[HC_7H_5O_2]}$$

$$6.3 \times 10^{-5} = \frac{(0.40 + \cancel{x})x}{(0.50 - \cancel{x})}$$

$$x = 7.9 \times 10^{-5}$$

$[H_3O^+] = 7.9 \times 10^{-5}$
pH = -log[H_3O^+]
pH = 4.10

b)   If you add 0.600 mol of H_3O^+ in 3.0 L, the equilibrium will shift left, therefore $[HC_7H_5O_2]$ will increase to 0.70 M and $[C_7H_5O_2^-]$ will decrease to 0.20 M

$HC_7H_5O_{2\ (aq)} + H_2O_{(l)} \rightleftharpoons C_7H_5O_2^-{}_{(aq)} + H_3O^+{}_{(aq)}$ $K_a = 6.3 \times 10^{-5}$
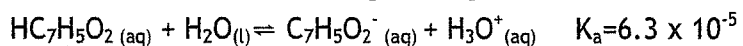
MR	1	1	1	1
I	0.70	---	0.20	0
C	-x	---	+x	+x
E	0.70-x	----	0.20+x	x

$0.20/K_a > 1000$

approx will work

$$K_a = \frac{[C_2H_3O_2^-][H_3O^+]}{[HC_2H_3O_2]}$$

$$6.3 \times 10^{-5} = \frac{(0.20 + \cancel{x})x}{(0.70 - \cancel{x})}$$

$$x = 2.2 \times 10^{-4}$$

$[H_3O^+] = 2.2 \times 10^{-4}$
pH = -log[H_3O^+]
pH = 3.66

c)   If you add 0.300 mol of OH^- in 3.0 L, the equilibrium will shift right, therefore $[HC_2H_3O_2]$ will decrease to 0.40 M and $[C_2H_3O_2^-]$ will increase to 0.50 M

$HC_7H_5O_{2\ (aq)} + H_2O_{(l)} \rightleftharpoons C_7H_5O_2^-{}_{(aq)} + H_3O^+{}_{(aq)}$ $K_a = 6.3 \times 10^{-5}$
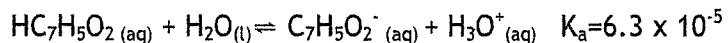
MR	1	1	1	1
I	0.40	---	0.50	0
C	-x	---	+x	+x
E	0.40-x	----	0.50+x	x

$0.40/K_a > 1000$

approx will work

$$K_a = \frac{[C_2H_3O_2^-][H_3O^+]}{[HC_2H_3O_2]}$$

$$6.3 \times 10^{-5} = \frac{(0.50 + \cancel{x})x}{(0.40 - \cancel{x})}$$

$$x = 5.0 \times 10^{-5}$$

$[H_3O^+] = 1.1 \times 10^{-5}$
pH = -log[H_3O^+]
pH = 4.30

7) a)
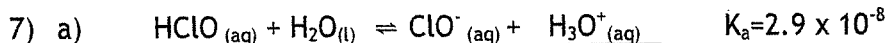
	$HClO_{(aq)}$	$+ H_2O_{(l)}$	$\rightleftharpoons ClO^-_{(aq)} +$	$H_3O^+_{(aq)}$
MR	1	1	1	1
I	0.45	---	0.35	0
C	-x	---	+x	+x
E	0.45-x	----	0.35+x	x

$K_a = 2.9 \times 10^{-8}$

$0.35/K_a > 1000$

approx will work

$$K_a = \frac{[ClO^-][H_3O^+]}{[HClO_2]}$$

$$2.9 \times 10^{-8} = \frac{(0.35 + \cancel{x})x}{(0.45 - \cancel{x})}$$

$$x = 3.7 \times 10^{-8}$$

$[H_3O^+] = 3.7 \times 10^{-8}$
$pH = -\log[H_3O^+]$
pH = 7.43

b) If you add 0.25 mol of H_3O^+ in 1.0 L, the equilibrium will shift left, therefore [HClO] will increase to 0.70 M and [ClO⁻] will decrease to 0.10 M
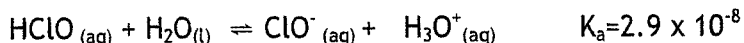
	$HClO_{(aq)}$	$+ H_2O_{(l)}$	$\rightleftharpoons ClO^-_{(aq)} +$	$H_3O^+_{(aq)}$
MR	1	1	1	1
I	0.70	---	0.10	0
C	-x	---	+x	+x
E	0.70-x	----	0.10+x	x

$K_a = 2.9 \times 10^{-8}$

$0.10/K_a > 1000$

approx will work

$$K_a = \frac{[ClO^-][H_3O^+]}{[HClO_2]}$$

$$2.9 \times 10^{-8} = \frac{(0.10 + \cancel{x})x}{(0.70 - \cancel{x})}$$

$$x = 2.0 \times 10^{-7}$$

$[H_3O^+] = 2.0 \times 10^{-7}$
$pH = -\log[H_3O^+]$
pH = 6.69

c) If you add 0.15 mol of OH⁻ in 1.0 L, the equilibrium will shift right, therefore [HClO] will decrease to 0.30 M and [ClO⁻] will increase to 0.50 M
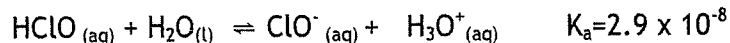
	$HClO_{(aq)}$	$+ H_2O_{(l)}$	$\rightleftharpoons ClO^-_{(aq)} +$	$H_3O^+_{(aq)}$
MR	1	1	1	1
I	0.30	---	0.50	0
C	-x	---	+x	+x
E	0.30-x	----	0.50+x	x

$K_a = 2.9 \times 10^{-8}$

$0.30/K_a > 1000$

approx will work

$$K_a = \frac{[ClO^-][H_3O^+]}{[HClO_2]}$$

$$2.9 \times 10^{-8} = \frac{(0.50 + \cancel{x})x}{(0.30 - \cancel{x})}$$

$$x = 1.7 \times 10^{-8}$$

$[H_3O^+] = 1.7 \times 10^{-8}$
$pH = -\log[H_3O^+]$
pH = 7.77