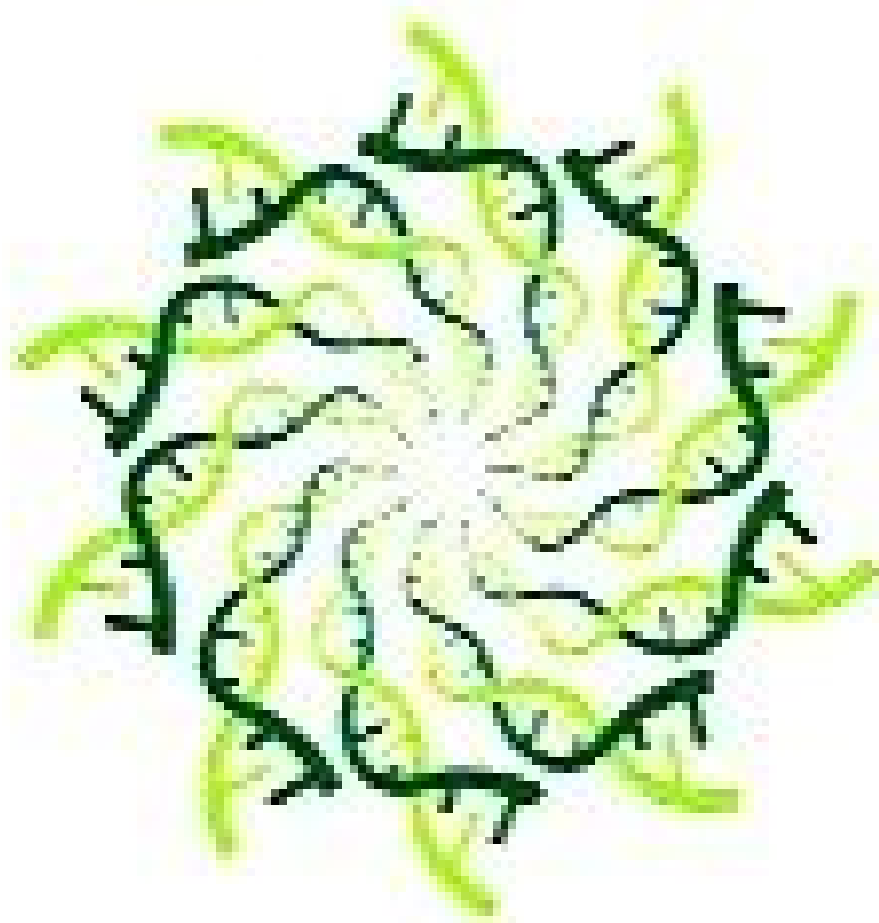






As above,  
so below...



# Acids & Bases Tutorial 2023.pdf

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## Tutorial 11: Sample Past Exam Questions/Acids and Bases

### QUESTION ONE

- (a) Distinguish between strong acid and weak acid and between concentrated acid and dilute acid. Illustrate your answers. [4]
- (b) Define each of the following
- (i) Arrhenius Acid (ii) Bronsted-Lowry Acid (iii) Lewis Acid [3]
- (c) Pyridine ( $C_5H_5N$ ) is an organic base ( $K_b = 1.7 \times 10^{-9}$ ).
- (i) Write its acid-base reaction in water and clearly show the conjugate base and conjugate acid of this reaction. [2]
- (ii) Write its corresponding  $K_b$  equilibrium expression. [2]
- (iii) The base and conjugate base compete for the proton in pyridine's acid-base reaction. Predict what happens to the position of equilibrium if the conjugate base has a higher affinity to the proton than pyridine. [1]
- (d) Calculate the pH of a 1.0 M solution of methylamine ( $CH_3NH_2$ ,  $K_b = 4.38 \times 10^{-4}$ ). [4]

### QUESTION TWO

- (a) Define each of the following using Arrhenius Model. [2]
- (i) Strong acid [½] (ii) Weak acid [½] (iii) Strong base [½] (iv) Weak base [½]
- (b) Define each of the following
- (i) Arrhenius Acid (ii) Bronsted-Lowry Acid (iii) Lewis Acid [3]
- (c) Which of the definitions in (b) above is the most general? Write reactions to justify your answer. [3]
- (d) Define what is meant by a conjugate acid-base pair. For each of the following reactions identify the acid, the base, conjugate acid and conjugate base [4]
- (i)  $H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$
- (ii)  $CH_3O^- + CH_3CCH_3 \rightleftharpoons CH_3OH + CH_3C=CH_2$
- (iii)  $H_2S + NH_3 \rightleftharpoons HS^- + NH_4^+$
- (iv)  $C_6H_5N + H_2O \rightleftharpoons C_6H_5NH^- + OH^-$
- (e) Define pH and pOH. Calculate the pH of each solution given below. [2]
- (i)  $[H^+] = 1.4 \times 10^{-4} M$
- (ii)  $[H^+] = 2.5 \times 10^{-10} M$
- (iii)  $[OH^-] = 3.5 \times 10^{-2} M$
- (iv)  $pOH = 2.3$
- (f) What are the major species in a 0.250 solution of perchloric acid ( $HClO_4$ )? Calculate its pH. [2]
- (g) For trichlorophenol ( $HC_6H_2Cl_3O$ )  $K_a = 1.0 \times 10^{-6}$ . Calculate the concentration of all species and the pH of a 0.05 solution of trichlorophenol in water. [5]



Rotate



Search



Share

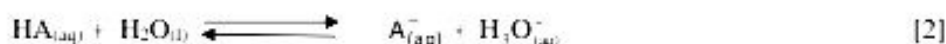
## Tutorial 11: Sample Past Exam Questions/Acids and Bases

### 3. QUESTION THREE

- (a) The **Bronsted-Lowry** definition of acids-bases is more general compared to the Arrhenius concept. Identify at least **2 limitations** of the Arrhenius definition of acids and bases?

[2]

- (b) What is a **conjugate acid-base pair** and use the equation below to identify a **conjugate base** and **base**



- (c) Clarify the difference between **weak acid** and **dilute acid**

[2]

- (d) Indicate whether solutions with each of the following ion concentrations are neutral, acidic or basic:

(i)  $[\text{H}^{+}] = 4.0 \times 10^{-9} \text{ M}$

(ii)  $[\text{OH}^{-}] = 1.0 \times 10^{-7} \text{ M}$

(iii)  $[\text{OH}^{-}] = 1.0 \times 10^{-13} \text{ M}$  [3]

- (e) A student prepared a 0.03 M solution of formic acid ( $\text{HCHO}_2$ ) and measured its pH at 25°C, and it was found to be 4.67

(i) Calculate  $K_a$  for formic acid at this temperature

(ii) What percentage of the acid is ionized in this 0.03 M solution? [3]

- (f) A chemist prepared a 0.010 M  $\text{NH}_3$  solution and by a freezing-point-lowering experiment determined that the  $\text{NH}_3$  had undergone 15.98% ionization. Calculate the  $K_b$  for  $\text{NH}_3$

[4]

- (g) What are the concentrations of all species present in a 0.5 M ethanoic acid [ $\text{H}_3\text{CCOOH}$ ] solution, given that  $K_a$  for  $\text{H}_3\text{CCOOH}$  is  $1.8 \times 10^{-5}$  [4]

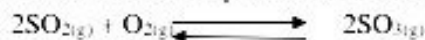
### 4. QUESTION FOUR

- (a) For the equilibrium



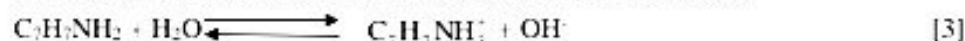
At 275°C,  $K_p$  is  $1.14 \times 10^2 \text{ atm}^2$  calculate the value of  $K$  and determine the values of  $K$  and  $K_p$  for the reverse reaction at 275°C [6]

- (b) A mixture of 0.0080 mol  $\text{SO}_{2(\text{g})}$  and 0.0056 mole of  $\text{O}_{2(\text{g})}$  is placed in a one liter container at 1000K. When equilibrium is established, 0.0040 mol  $\text{SO}_{3(\text{g})}$  is present



What are the equilibrium concentrations of  $\text{SO}_{2(\text{g})}$  and  $\text{O}_{2(\text{g})}$  and determine the value of  $K_p$  for the equilibrium at 1000K [6]

- (c) In a 0.25 M solution of benzyl amine ( $\text{C}_7\text{H}_7\text{NH}_2$ ) the concentration of  $\text{OH}^{-}$  is  $2.4 \times 10^{-3} \text{ M}$ . What is the ionization constant for this weak base, the reaction is



- (d) What are the limitations of the Bronstead-Lowry concept? [2]

- (e) Calculate the pH of a buffered solution containing 0.25 M acetic acid and 0.10 M sodium acetate. [3]



# ACIDS, BASES, AND SALTS

1(a) Strength of an acid is determined by the equilibrium position.

Strong acid is one whose equilibrium lies far to the right while a weak acid is one whose equilibrium lies far to the left.

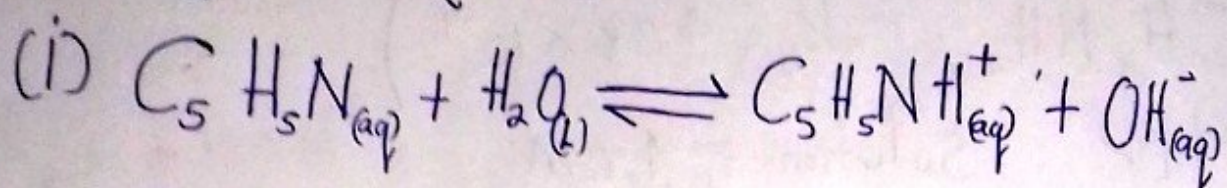
A concentrated acid is one which contains very large amount of acid than  $H_2O$  in the solution while a dilute one has greater concentrations of  $H_2O$  in it.

(b) Just give them from the notes

- (i) Produces  $H^+$  when dissolved in  $H_2O$ .
- (ii) Proton donor (Bronsted-Lowry acid)
- (iii) Electron Pair Acceptor. (Lewis Acid)



(c) Pyridine ( $C_5H_5N$ )  $K_b = 1.7 \times 10^{-9}$ .



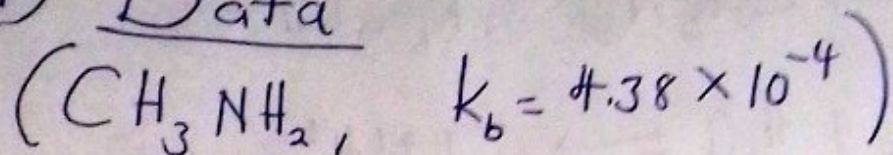
Conjugate acid =  $C_5H_5NH^+$

Conjugate base =  $OH^-$

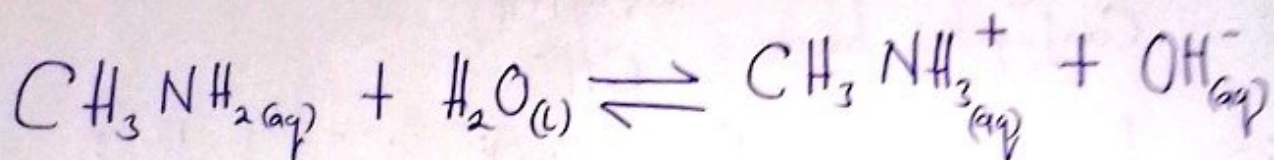
$$(ii) K_b = \frac{[OH^-][C_5H_5NH^+]}{[C_5H_5N]}$$

(iii) If the Conjugate base has a higher affinity for  $H^+$  (a proton) than does Pyridine, the equilibrium position will be far to the left.

(d) Data



PH=? Solution = 1.0 M



$$K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]}$$

$$4.38 \times 10^{-4} = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]}$$

We write the ICE table

	$\text{CH}_3\text{NH}_{2(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{CH}_3\text{NH}_3^+_{(aq)} + \text{OH}^-_{(aq)}$			
Initial	1.0		0	0
$\Delta$ (change) in Concentration	-x		+x	+x
Equilibrium	1.0 - x		x	x



Back to our  $K_b$  expression

$$K_b = 4.38 \times 10^{-4} = \frac{[\text{OH}^-][\text{CH}_3\text{NH}_3^+]}{[\text{CH}_3\text{NH}_2]}$$
$$= \frac{x \times x}{(1.0 - x)} = \frac{x^2}{1.0 - x}$$

We assume that  $1.0 - x = 1$  (1 is much much greater than  $x$  such that  $1.0 - x = 1$ )

$$\therefore 4.38 \times 10^{-4} = \frac{x^2}{1}$$

$$x = \sqrt{4.38 \times 10^{-4}} = \underline{\underline{0.02 \text{ M}}}$$

$$\therefore x = [\text{OH}^-] = 0.02 \text{ M}$$

$$\text{and } \text{pOH} = -\log[\text{OH}^-]$$
$$= -\log[0.02]$$
$$= 1.70$$

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

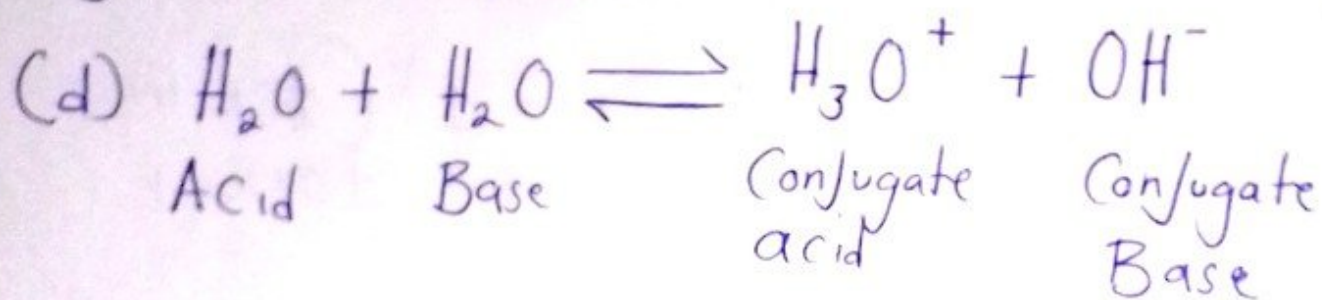
$$\text{pH} = 14 - 1.70$$
$$= \underline{\underline{12.3}}$$



Nothing ventured  
Nothing gained 🌀❄️

## QUESTION 2

We all know the definitions So im gonna skip them. Just refer to you notes!



Let's go to (F) and (g)

Men!

For (e) we just manipulate 3 Formulas . .

$$\text{pH} = -\log[\text{H}^+]$$

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

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

(F) We have 0.250M solution of  $(\text{HClO}_4)$

Since Perchloric acid is a strong acid,  
It will ionize completely in  $\text{H}_2\text{O}$ .

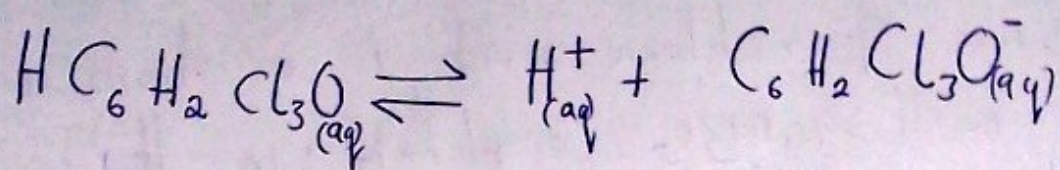
Major species are  $(\text{H}^+)$  and  $(\text{ClO}_4^-)$  and  $\text{H}_2\text{O}$ .

Since the concentration of  $\text{HClO}_4$  is  
0.250M, the concentration of  $\text{H}^+$  ions is  
also 0.250M.

$$\begin{aligned}\text{pH} &= -\log [\text{H}^+] \\ &= -\log (0.250\text{M}) \\ &= \underline{\underline{0.602}}\end{aligned}$$



(9) We have trichlorophenol with  $K_a = 1 \times 10^{-6}$



$$\text{We have } K_a = \frac{[\text{H}^+][\text{C}_6\text{H}_2\text{Cl}_3\text{O}^-]}{[\text{HC}_6\text{H}_2\text{Cl}_3\text{O}]}$$

We write the ICE Table:

	$\text{HC}_6\text{H}_2\text{Cl}_3\text{O}$	$\rightleftharpoons$	$\text{H}_{(aq)}^+$	+	$\text{C}_6\text{H}_2\text{Cl}_3\text{O}_{(aq)}^-$
Initial	0.05M		0		0
$\Delta$ concs	-x		+x		+x
Equilibrium	0.05-x		x		x

$$1 \times 10^{-6} = \frac{x \times x}{0.05 - x}$$

We assume  $0.05 \gg x$  such that  
 $0.05 - x = 0.05$



$$1 \times 10^{-6} = \frac{x^2}{0.05}$$

$$\sqrt{x^2} = \sqrt{5 \times 10^{-8}}$$

$$x = \underline{\underline{2.24 \times 10^{-4} \text{ M}}}$$

So our  
Concentrations for  
the Species are

$$[\text{HC}_6\text{H}_2\text{Cl}_3\text{O}] = 0.05 - x = 0.05 - 2.24 \times 10^{-4} = \underline{\underline{0.05}}$$

$$[\text{H}^+] = x = \underline{\underline{2.24 \times 10^{-4} \text{ M}}}$$

$$\text{C}_6\text{H}_2\text{Cl}_3\text{O}^- = x = \underline{\underline{2.24 \times 10^{-4} \text{ M}}}$$

$$\text{pH} = -\log [\text{H}^+]$$

$$= -\log [2.24 \times 10^{-4}]$$

$$= \underline{\underline{3.65}}$$

### Validity test

$$\frac{x}{H_A} \times 100 < 5\%$$

$$\frac{2.24 \times 10^{-4}}{0.05} \times 100 = 0.448$$

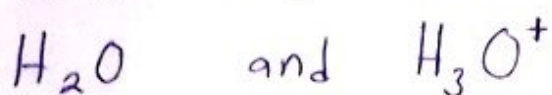
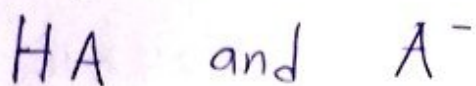
Since 0.448 is less  
than 5% then  
our approximation  
is considered.



## QUESTION 3

- (a) 2 limitations of Arrhenius definition
- It applies to the aqueous solutions
  - Allows for only one kind of base which is the hydroxide ion ( $\text{OH}^-$ ).

(b) Conjugate acid-base pair is something that consists of 2 substances related to each other by donating and accepting a proton.



(c) Check Question 1 solutions

(d) (i)  $[\text{H}^+] = 4.0 \times 10^{-9} \text{ M}$

$$\text{pH} = -\log [\text{H}^+] = -\log [4.0 \times 10^{-9} \text{ M}]$$

$$= \underline{\underline{8.40}} \quad \text{hence solution is basic!}$$

(ii)  $[\text{OH}^-] = 1 \times 10^{-7} \text{ M}$

$$\text{pOH} = -\log [\text{OH}^-] = -\log [1 \times 10^{-7}]$$

$$= 7$$

$$\text{pH} = 14 - 7 = \underline{\underline{7}}$$

Hence solution is neutral!



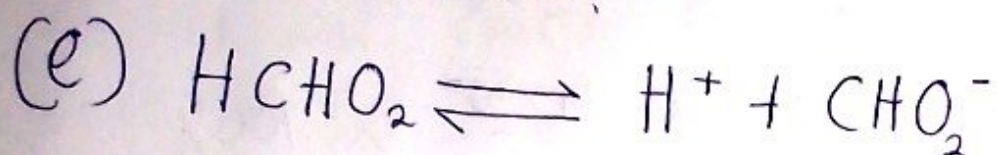
$$(iii) [OH^-] = 1.0 \times 10^{-13} M$$

$$pOH = -\log [OH^-] = -\log [1.0 \times 10^{-13}]$$

$$= 13$$

$$pH + pOH = 14$$

$$pH = 14 - 13 = \underline{\underline{1}} \quad \text{Hence solution is acidic,}$$



$$pH = 4.67 \quad \text{and} \quad pH = -\log [H^+]$$

$$[H^+] = 10^{-pH} = 10^{-4.67} = \underline{\underline{2.14 \times 10^{-5} M}}$$

? We write the ICE table (oh no!)

✓ Let's write the  $K_a$  expression  
Fellas

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

$$\text{But } [H^+] = [CHO_2^-] = 2.14 \times 10^{-5} M$$

and we know that equilibrium concentration for  $HCHO_2$  is  $0.03M - x$ ,  
we can assume that  $0.03M - x = 0.03M$



## the spiral

One of our most ancient  
and enduring symbols,  
it represents the winding journeys  
we must take inward if we're to  
truly know & love ourselves.  
From these never-ending journeys  
we return with infinitely more  
power and wisdom.



$$\text{So } K_a = \frac{(2.14 \times 10^{-5})(2.14 \times 10^{-5})}{0.03}$$

$$\therefore K_a = \underline{\underline{1.53 \times 10^{-8}}}$$

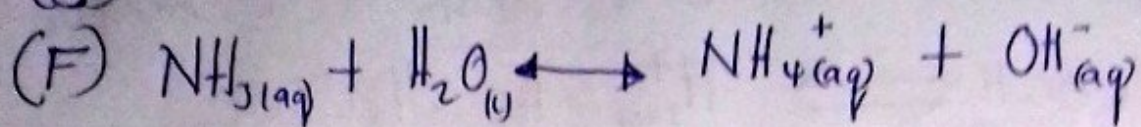
$$(ii) \% \text{ ionization} = \frac{\text{Amount ionized}}{\text{Initial concs}} \times 100\%$$

$$= \frac{2.14 \times 10^{-5}}{0.03} \times 100\%$$

$$= \underline{\underline{0.07\%}}$$



(E) Concentrations of  $\text{NH}_3 = 0.010\text{M}$



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

Equilibrium concs for all the species will be  $\text{NH}_3 = 0.010\text{M} - x$ ,

$$[\text{NH}_4^+] = [\text{OH}^-] = x$$

$$\text{therefore } \% \text{ ionization} = \frac{\text{Amount ionized}}{\text{Initial concs}} \times 100\%$$

Initially we had  $0.010\text{M}$ , Amount ionized is equal to  $[\text{OH}^-] = x$

$$\therefore \% \text{ ionization} = \frac{x}{0.010} \times 100\%$$

$$15.98\% = \frac{x}{0.010} \times 100\%$$

$$15.98 = 10000x$$

$$x = 1.598 \times 10^{-3}$$

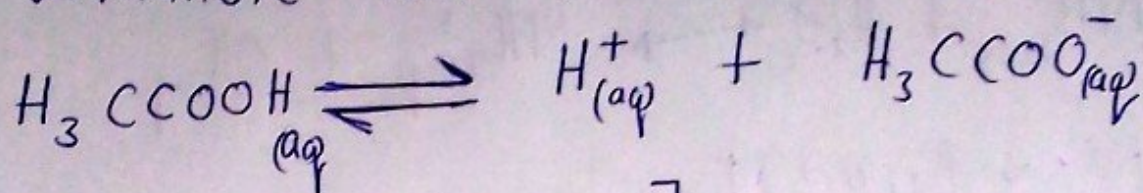
then  $K_b = \frac{[1.598 \times 10^{-3}]^2}{0.010}$

$$= 2.55 \times 10^{-4}$$

                  ||



(9) Ethanoic acid is a weak acid therefore



$$K_a = \frac{[\text{H}^+][\text{H}_3\text{CCOO}^-]}{[\text{H}_3\text{CCOOH}]}$$

ICE table men!

	$\text{H}_3\text{CCOOH}$	$\rightleftharpoons$	$\text{H}^+$	$\text{H}_3\text{CCOO}^-$
Initial concs	0.5 M		0	0
$\Delta$ concs	$-x$		$+x$	$+x$
Equilibrium	$0.5 - x$		$x$	$x$

$$K_a = \frac{[\text{H}^+][\text{H}_3\text{CCOO}^-]}{[\text{H}_3\text{CCOOH}]}$$

$$[\text{H}^+] = x$$

$$[\text{H}_3\text{CCOO}^-] = x$$

$$[\text{H}_3\text{CCOOH}] = 0.5 - x$$

$$K_a = \frac{x \times x}{0.5 - x}$$

W/e assume that  $0.5 \gg x$  Such that

$$0.5 - x = 0.5$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.5}$$

$$\sqrt{x^2} = \sqrt{9 \times 10^{-6}} \quad x = \underline{\underline{3 \times 10^{-3} M}}$$

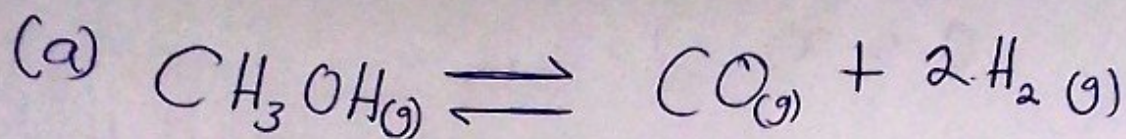
$$\text{then } [H^+] = x = \underline{\underline{3 \times 10^{-3} M}}$$

$$[H_3CCOO^-] = x = \underline{\underline{3 \times 10^{-3} M}}$$

$$[H_3CCOO] = 0.5 - 3 \times 10^{-3} M = \underline{\underline{0.5 M}}$$



## QUESTION 4



We know that  $K_p = K_c (RT)^{\Delta n}$

therefore

$$1.14 \times 10^2 = K_c (0.08206 \times (273.15 + 275))^{\Delta n}$$

$$114 = K_c (44.98)^{\Delta n}$$

$$\Delta n = (2+1) - (1) = 3-1 = \underline{2}$$

$$114 = K_c (44.98)^2$$

$$K_c = 0.056 \approx \underline{\underline{0.06}}$$

$$K_{c \text{ reverse}} = \frac{1}{0.06} = \underline{\underline{16.67}}$$

$$K_{p \text{ reverse}} = \frac{1}{114} = \underline{\underline{8.77 \times 10^{-3}}}$$

(C) Benzyl Amine ( $C_7H_7NH_2$ )

$$[OH^-] = 2.4 \times 10^{-3} M.$$

$$\text{Solution} = 0.25 M.$$

⚠ Short cut to solve this problem

$$[OH^-] = \sqrt{K_b [\text{Base}_{\text{solution}}]}$$

$$2.4 \times 10^{-3} = \sqrt{K_b (0.25)}$$

$$K_b = \frac{(2.4 \times 10^{-3})^2}{0.25}$$

$$K_b = \underline{\underline{2.30 \times 10^{-5}}}$$



(d) - It fails to explain the reaction between acid oxides and basic oxides because there is no proton transfer.

(e) Here we will use the Henderson-Hasselbalch equation.

$$pH = pK_a + \log \frac{[Base]}{[Acid]}$$

$$pK_a = -\log K_a$$

But to calculate the  $pK_a$  we need  $K_a$ . So after conducting some research the direct  $pK_a$  value for acetic acid is 4.75. (meaning it was supposed to be given in the question)

$$pH = 4.75 + \log \frac{(0.10)}{(0.25)}$$

$$= 4.75 - 0.398$$

$$= \underline{\underline{4.35}}$$



GRADE	RANGE	CBU GRADE POINT EQUIVALENT	
		FULL COURSE	HALF COURSE
<b>A<sup>+</sup></b>	<b>≥ 86%</b>	<b>5</b>	<b>2.5</b>
<b>A</b>	<b>76% - 85%</b>	<b>4</b>	<b>2</b>
<b>B<sup>+</sup></b>	<b>68% - 75%</b>	<b>3</b>	<b>1.5</b>
<b>B</b>	<b>62% - 67%</b>	<b>2</b>	<b>1</b>
<b>C<sup>+</sup></b>	<b>56% - 61%</b>	<b>1</b>	<b>0.5</b>
<b>C</b>	<b>50% - 55%</b>	<b>0</b>	<b>0</b>
<b>D<sup>+</sup></b>	<b>40% - 49%</b>	<b>0</b>	<b>0</b>
<b>D</b>	<b>≤39%</b>	<b>0</b>	<b>0</b>

*Where there are 12 courses*

Distinction : 42.5 and above

Merit : 30.5 - 42

Credit : 18.5 - 30

Pass : 0 – 18





