CHAPTER 7 : IONIC EQUILIBRIA

SESI 2010/2011

1. (a) (i) What is meant by pH of solution? [1 mark]

   (ii) Derive the relationship between the pH and pOH of a solution. [2 marks]

   (b) Pyridine, C_5H_5N has a pK_b = 8.76 , is a bad-smelling liquid.

   (i) Calculate Kb for C_5H_5N. [2 marks]

   (ii) Calculate pH of 0.20 M pyridine. [5 marks]

2. (a) Define the equivalence point and the end point for a titration. [2 marks]

   (b) A 30.0 mL HCl solution is titrated to the end point by 20.0 mL of 0.2 M NaOH solution. Calculate the molarity of the HCl solution. [3 marks]

   (c) At 25^0C ,2.20 % of benzoic acid, C_6H_5COOH in 0.125 M solution is ionized. Write the ionisation equation for benzoic acid in water. Determine the acid dissociation constant, Ka for benzoic acid and calculate pH for the solution. [5 marks]

NO	PART	SCHEME	MARK
1	(a)(i)	pH is define as the negative of the logarithm of $[H_3O^+]$ @ $[H^+]$ pH = -log $[H_3O^+]$	1
	(a)(ii)	$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ $-\log K_w = -\log [H_3O^+][OH^-] = -\log 1.0 \times 10^{-14}$ $pK_w = -\log [H_3O^+] - \log [OH^-] = 14$ pK_w = pH +pOH = 14 @ pH = pK_w – pOH @ = 14 -pOH	
	(b)(i)	K_b = antilog ($-pK_b$) @ = antilog (- 8.76) = 1.74×10^{-9}	
	(b)(ii)	pH of 0.2 M $C_5H_5N + H_2O \rightleftharpoons C_5H_5NH^+ + OH^-$ $K_b = \frac{[C_5H_5NH^+][OH^-]}{[C_5H_5N]- x}$	

		Since K_b << 1,assume that $C_5H_5N - x \approx 0.2$ $1.74\ x\ 10^{-9} = \frac{x^2}{0.2}$ X = $[OH^-] = 1.865\ x\ 10^{-5}$ M pOH = -log $[OH^-]$ = -log $(1.865\ x\ 10^{-5})$ = 4.73 pH = 14 – 4.73 = 9.27	
2	(a)	Equivalence point The point in a titration the chemical amount of titrant added is equal to the chemical amount of the substance being titrated @ mol H^+ = mol OH^- End point The point in a titration at which the indicator signals that a stoichiometric amount of the first reactant has been added to the second reactant. @ colour of indicator changes when its reach equivalence point.	
	(b)	$HCl + NaOH \longrightarrow NaCl + H_2O$ @ nHCl = nNaOH nHCl = $(20.0 \times 10^{-3}) \times 0.2$ = 4.0×10^{-3} mol Molarity HCl = $\frac{4.0\ x\ 10^{-3}}{30.0\ x\ 10^{-3}}$ = 0.13 $molL^{-1}$	
	(c)	$C_6H_5COOH\ (aq) \rightleftharpoons C_6H_5COO^-\ (aq) + H^+\ (aq)$ $[\ ]_0$ 0.125 0 0 $[\ ]_\Delta$ -x +x +x $[\ ]_{eq}$ 0.125-x x x Ka = $\frac{[C_6H_5COO^-][H^+]}{[C_6H_5COOH]}$ Ka = $\frac{x^2}{0.125-x}$ α = 2.2 % α = $\frac{[H^+]}{[C_6H_5COOH]}$ X 100%	

		$[H^+] = \frac{2.1\ x\ 0.125}{100}$ $= 2.75 \times 10^{-3}$ M X = 6.186×10^{-5} OR Assume $0.125 - 2.75 \times 10^{-3} \approx 0.125$ $= \frac{(2.75\ x\ 10^{-3})^2}{0.125}$ $= 6.05 \times 10^{-5}$ pH = $-\log(2.75 \times 10^{-3})$ = 2.56	

SESI 2011/2012

1. (a) Define Bronsted-Lowry acid and base. [2 marks]

(b) Explain how 100mL HCl solution with pH 1.05 can be prepared from 8.50 M HCl. [4 marks]

(c) In an acid-base titration, 10 mL of 0.50 M HCl was added to 40 mL of 0.10 M NaOH. Determine the pH of the solution formed. [4 marks]

2. (a) At 25^0C, 0.69 % hydrazine is ionized in 0.02 M hydrazine solution.

(i) Calculate the concentration of OH^- ion in the solution. [3 marks]

(ii) Calculate the ionisation constant, Kb of hydrazine. [2 marks]

(b) Ionisation reaction of phenylacetic acid, $C_6H_5CH_2COOH$ is as follows:

$$C_6H_5CH_2COOH + H_2O \rightleftharpoons C_6H_5CH_2COO^- + H_3O^+$$

(i) Calculate the concentration of $C_6H_5CH_2COO^-$ ion in 0.19 M solution of $C_6H_5CH_2COOH$. (Kb = 4.90×10^{-5}) [4 marks]

(ii) What is the pH of this solution? [1 marks]

NO	PART	SCHEME	MARKS
1	(a)	An acid is a proton donor A base is a proton acceptor	1 1
	(b)	pH = $-\log[H^+]$ = 1.05 $[H^+]$ = 0.089 M $M_1V_1 = M_2V_2$ @	1

		$\frac{(0.089)(100)}{8.5} = V_1$ V_1 = 1.05 mL 1.05 mL HCl 8.5 M is added with distilled water until the volume becomes 100 mL.	1 1 1
	(c)	HCl + NaOH $\longrightarrow$ NaCl + H_2O $n_i = \frac{10 \, x \, 0.5}{1000}$ $\frac{40 \, x \, 0.1}{1000}$ @ $= 5 \, x \, 10^{-3}$ $4 \, x \, 10^{-3}$ $n_f = 1 \, x \, 10^{-3}$ 0 $4 \, x \, 10^{-3}$ 0 $[H^+] = \frac{1 \, x \, 10^{-3}}{\frac{5}{1000}}$ = 0.02 M pH = -log [0.02] = 1.7	 1 1 1 1
2	(a) (i) (a)(ii)	$N_2H_4 + H_2O \rightleftharpoons N_2H_5^+ + OH^-$ $[OH^-] = [N_2H_5^+]$ = (0.69 /100) x 0.020 = 1.38 x 10^{-4} M Kb = $[N_2H_5^+] [OH^-] / [N_2H_4]$ $= \frac{(1.38 \, x \, 10^{-4})^2}{0.020 - (1.38 \, x \, 10^{-4})}$ = 9.59 x 10^{-7}	1 1 1 1 1
	(b)(i)	$C_6H_5CH_2COOH + H_2O \rightleftharpoons C_6H_5CH_2COO^- + H_30^+$ [ ] 0.19 M 0 0 Δ -x +x +x Eq. 0.19-x x x Ka = $\frac{[C_6H_5CH_2COO^-][H_30^+]}{[C_6H_5CH_2COOH]}$ @	 1 1

		$= \frac{x^2}{0.19 - x}$	1
		Since Ka << 1, assume that 0.19 – x ≈ 0.19	
		$4.9 \times 10^{-5} = \frac{x^2}{0.19 - x}$	1
		X = $3.05\ x\ 10^{-3}$	1
		X = $[C_6H_5CH_2COO^-]$ = $[H_30^+]$ = $3.05\ x\ 10^{-3}$ M	
	(b)(ii)	pH = -log $[H_3O^+]$ = -log (= $3.05\ x\ 10^{-3}$)	1
		= 2.52	

UPS TK025-SESI 2012/2013

1. (a) The pH of a fruit juice is 3.52. Calculate the concentration of $H^+(aq)$ ions present in the fruit juice. [2 marks]

   (b) Pyridine, C_5H_5N is a weak base which was discovered in coal tar in 1846. If the percentage dissociation of 0.0015 moldm^{-3} C_5H_5N is 0.10 %, calculate

   (i) the concentrations of OH^- at equilibrium

   (ii) the base dissociation constant, Kb for C_5H_5N

   (iii) the acid dissociation constant, Ka for its conjugate acid, $C_5H_5NH^+$

   [8 marks]

2. (a) Sodium benzoate, C_6H_5COONa is a salt formed when sodium hydroxide, NaOH reacts with benzoic acid, C_6H_5COOH [Ka= 6.3×10^{-5}]

   (i) Write an equation for this reaction.

   (ii) Classify the salt formed. Explain by using the appropriate equations.

   [4 marks]

   (b) Determine the pH of the solution formed when 30 mL of 0.25 m NaOH solution is titrated with 40 mL of 0.10 M of HCl solution.

   [6 marks]

NO	PART	SCHEME	MARK
1	(a)	pH = - log $[H^+]$ @ -log $[H^+]$ = 3.52 **$[H^+]$ = 3.0 x 10^{-4} M**	1 1
	(b) (i)	<table><tr><td>Equation</td><td>C_5H_5N (aq) + H_2O (l) $\rightleftharpoons$ $C_5H_5NH^+$ (aq)+ OH^-(aq)</td><td></td><td></td></tr><tr><td>[ ]initial</td><td>0.0015</td><td>0</td><td>0</td></tr><tr><td>[]change</td><td>-x</td><td>+x</td><td>+x</td></tr><tr><td>[ ]eq</td><td>0.0015-x</td><td>x</td><td>x</td></tr></table> α = $\frac{[C_5H_5\,N]change}{[C_5H_5N]initial}$ X 100 % @ $[C_5H_5N]_{\text{change}}$ = $\frac{0.10}{100}$ X 0.0015 = 1.5 x 10^{-6} moldm^{-3} = x $[OH^-]$ = 1.5 x 10^{-6} moldm^{-3}	1 1 1 1
	(ii)	$K_b = \frac{[C_5H_5NH^+][OH^-]}{[C_5H_5]}$ @ $= \frac{(1.5\ x\ 10^{-6})^2}{0.0015-x}$ = 1.5 x 10^{-9} moldm^{-3}	1 1
	(iii)	$K_a = \frac{K_w}{K_b} = \frac{1.0\ x\ 10^{-14}}{1.5\ x\ 10^{-9}}$ = 6.7 x 10^{-5} mol dm^{-3}	1 1
		TOTAL	**10**

NO	PART	ANSWER SCHEME	MARKS
2	(a)(i)	$C_6H_6COOH\ (aq) + NaOH\ (aq) \longrightarrow C_6H_5COONa\ (aq) + H_2O\ (l)$	1
	(ii)	**<u>Basic salt</u>**	1
		$C_6H_5COONa \longrightarrow C_6H_5COO^- + Na^+$	1
		$C_6H_5COO^- + H_2O \rightleftharpoons C_6H_6COOH + OH^-$	1
		Anion from the salt will undergo hydrolysis to give basic solution @ pH > 7	
	(b)	n(NaOH) = MV =(0.25)(0.03) = 7.5×10^{-3} mol	1
		n(HCl) = MV = (0.10)(0.04) = 4.0×10^{-3} mol	1
		(see table below)	1
		$M = \frac{3.5 \times 10^{-3}}{0.07}$ = **0.05 M**	1
		pOH = - log [OH^-] = - log (0.05) = **1.30**	1
		pH = 14 – pOH = 14 – 1.30 = **12.7**	1
		TOTAL	**10**

	NaOH	+ HCl	$\longrightarrow$ NaCl +	H_2O
$n_{initial}$	7.5×10^{-3}	4.0×10^{-3}	0	0
n_{change}	-4.0×10^{-3}	-4.0×10^{-3}	$+4.0 \times 10^{-3}$	$+4.0 \times 10^{-3}$
n_{final}	3.5×10^{-3}	0	$+4.0 \times 10^{-3}$	$+4.0 \times 10^{-3}$

SESI 2009/2010-TS027

1. a) i) Define Bronsted-Lowry acids and bases.
   ii) Calculate the pH of 0.003 M of HNO_3 acid [3 marks]

   b) The molarity of aqueous ethanoic acid (CH_3COOH) is 1.12 M.
   [Ka = 1.8×10^{-5} M].
   Calculate the

   i) pKa
   ii) concentration oh H^+ ions
   iii) degree of dissociation of acid [7 marks]

ANSWER SCHEME- SESI 2009/2010-TS027

NO	PART	SCHEME	MARKS
1	(a) i.	Acids are proton donors while bases are proton acceptors.	1
	(a) ii.	$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$ Initial: 0.003 0 0 Final: 0 0.003 0.003 $[H_3O^+]$ = 0.003 M @ pH = - log $[H_3O^+]$ @ - log 0.003 = 2.52	 1 1
	(b)i	pKa = -log Ka = - log 1.8×10^{-5} = 4.74	1 1
	(b)ii.	$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$ Initial: 1.12 0 0 Final: 1.12-x x x $Ka = \frac{[CH_3COO^-][CH_3COO^-]}{[CH_3COOH]}$ @ $1.8 \times 10^{-5} = \frac{x^2}{1.12-x}$ Assume x is small, 1.12 – x ≈ 1.12	 1 1

		$X = [H_3O^+]$ $= 4.49\ x\ 10^{-3}$ M $\alpha = \frac{[\ \ ]change}{[\ \ ]initial}$ @ $= \frac{4.49\ x\ 10^{-3}}{1.12}$ $= 4.0 \times 10^{-3}$	1 1 1
		TOTAL	**10**