Topic 7.2 : Ionic Equilibrium

1 ACJC/2009/P3/Q4(a)

4(a)(i) $K_{sp} = [Mg^{2+}][OH^-]^2$

(ii) Solubility of $Mg(OH)_2 = 9.08 \times 10^{-3}$ g dm^{-3}

$= \frac{9.08 \times 10^{-3}}{58.3}$

$= 1.557 \times 10^{-4}$ mol dm^{-3}

∴ In a saturated solution, $[Mg^{2+}] = 1.557 \times 10^{-4}$ mol dm^{-3}

$[OH^-] = 2 \times 1.557 \times 10^{-4}$

$= 3.114 \times 10^{-4}$ mol dm^{-3}

Hence, $K_{sp}(Mg(OH)_2) = (1.557 \times 10^{-4})(3.114 \times 10^{-4})^2$

$= 1.510 \times 10^{-11}$ mol^3 dm^{-9}

$= 1.5 \times 10^{-11}$ mol^3 dm^{-9} (2 sf)

(iii) The solubility of $Mg(OH)_2$ is lower in $Mg(NO_3)_2$ (aq).

$$Mg(OH)_2\,(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$$

In $Mg(NO_3)_2$ (aq), $[Mg^{2+}]$ is higher. The equilibrium above thus shifts to the left to use up some of the excess Mg^{2+}.

(iv) $[Mg^{2+}] = \frac{50 \times 3.2 \times 10^{-4}}{80}$

$= 2.0 \times 10^{-4}$ mol dm^{-3}

$[OH^-] = \frac{30 \times 5.0 \times 10^{-4}}{80}$

$= 1.875 \times 10^{-4}$ mol dm^{-3}

Now, $[Mg^{2+}][OH^-]^2 = (2.0 \times 10^{-4})(1.875 \times 10^{-4})^2$

$= 7.03 \times 10^{-12}$ mol^3 dm^{-9}

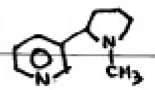
$< K_{sp}(Mg(OH)_2)$

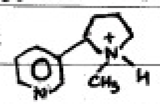
∴ No precipitation occurs.

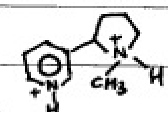
2 AJC/2009/P2/Q4(a)-(b)

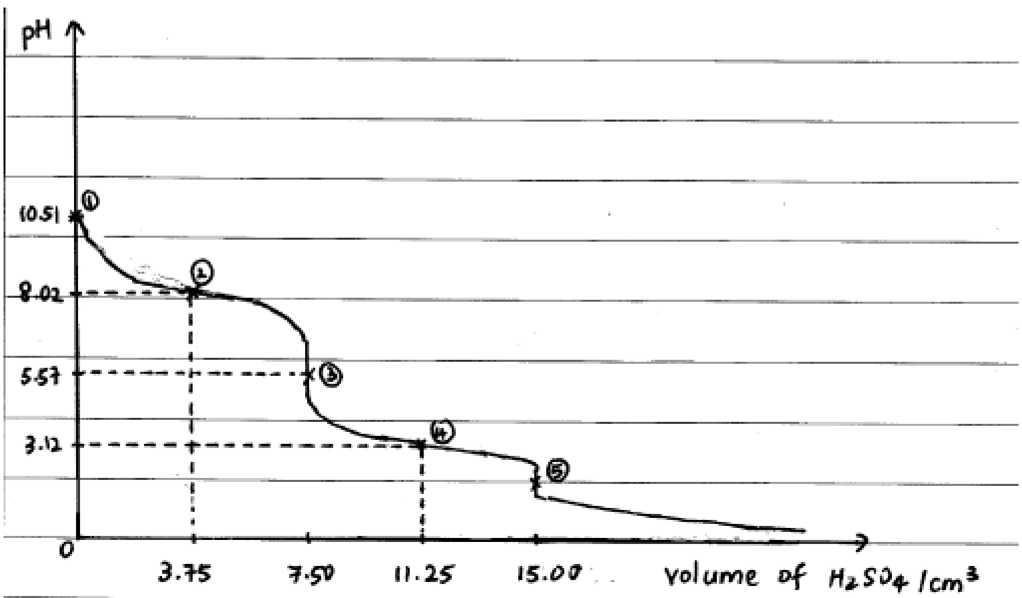
4(a) (Not required by syllabus — specially written for enrichment purposes only)

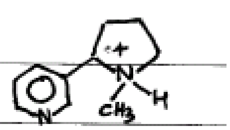
Point	$V_{H_2SO_4}$ / cm^3	Significance	Calculation of pH
①	0.00	Initial pH	$[OH^-] = \sqrt{cK_{b1}} = 3.236 \times 10^{-4}$ mol dm^{-3} pOH = 3.49 ⇒ pH = 10.51
②	3.75	max. buffer capacity (B_2/B_2H^+)	pOH = pK_{b1} = 5.98 ⇒ pH = 8.02
③	7.50	B_2 completely neutralised to B_2H^+	pOH = $\frac{pK_{b1} + pK_{b2}}{2}$ = 8.43 ⇒ pH = 5.57
④	11.25	max. buffer capacity ($B_2H^+/B_2H_2^{2+}$)	pOH = pK_{b2} = 10.88 ⇒ pH = 3.12
⑤	15.00	B_2H^+ completely neutralised to $B_2H_2^{2+}$	pH ≈ 2 (very hard to calculate exactly)

B_2 = 

B_2H^+ = 

$B_2H_2^{2+}$ = 



(b) 

3 AJC/2009/P3/Q4(a)-(b), (e)

4(a) $CH_3CH(OH)CO_2H \rightleftharpoons CH_3CH(OH)CO_2^- + H^+$

At equilibrium, $[H^+] = [CH_3CH(OH)CO_2^-] = 10^{-6.50}$ mol dm^{-3}

$[CH_3CH(OH)CO_2H] = (x - 10^{-6.50})$ mol dm^{-3}

Now, $K_a = \dfrac{[H^+][CH_3CH(OH)CO_2^-]}{[CH_3CH(OH)CO_2H]} \Rightarrow 10^{-3.86} = \dfrac{(10^{-6.50})^2}{x - 10^{-6.50}}$

$\therefore x = 3.17 \times 10^{-7}$ mol dm^{-3}

(b)(i) $n_{CH_3CH(OH)CO_2^-Na^+} = \dfrac{22.4}{112.1} = 0.200$ mol

$\therefore [CH_3CH(OH)CO_2^-] = 0.200$ mol dm^{-3}

Now, $K_a = \dfrac{[H^+][CH_3CH(OH)CO_2^-]}{[CH_3CH(OH)CO_2H]} \Rightarrow 10^{-3.86} = \dfrac{[H^+] \times 0.20}{0.10}$

$\therefore [H^+] = 6.90 \times 10^{-5}$ mol dm^{-3}

pH = 4.16

(ii) $CH_3CH(OH)CO_2^- + H^+ \rightarrow CH_3CH(OH)CO_2H$

$n_{CH_3CH(OH)CO_2^-}$ remaining $= 0.200 - \dfrac{10.0}{1000} \times 0.10$

$= 0.199$ mol

$n_{CH_3CH(OH)CO_2H} = 0.10 + \dfrac{10.0}{1000} \times 0.10$

$= 0.101$ mol

$\therefore [CH_3CH(OH)CO_2^-] = \dfrac{0.199}{1.01} = 0.197$ mol dm^{-3}

$[CH_3CH(OH)CO_2H] = \dfrac{0.101}{1.01} = 0.100$ mol dm^{-3}

Now, $10^{-3.86} = \dfrac{[H^+] \times 0.197}{0.100}$

$\therefore [H^+] = 7.01 \times 10^{-5}$ mol dm^{-3}

pH = 4.15

(e) (i) K_{sp} $(Fe(OH)_2)$ = $[Fe^{2+}][OH^-]^2$

For precipitation to occur, $[Fe^{2+}][OH^-]^2 > K_{sp}$

$(0.010) \times [OH^-]^2 > 1.6 \times 10^{-14}$

$[OH^-] > 1.265 \times 10^{-6}$ mol dm^{-3}

$\therefore$ pOH = $-\lg(1.265 \times 10^{-6}) = 5.90$

pH required = 14 − 5.90 = 8.10

4 CJC/2009/P3/Q3(c)

(c)

(i) Since K_{stab} of $[Fe(CN)_6]^{3-}$ is larger than $[Fe(SCN)(H_2O)_5]^{2+}$, $[Fe(CN)_6]^{3-}$ is the more stable complex. Ligand exchange will occur.
Deep red solution turns orange yellow.

(ii) $Fe^{3+} + e \rightarrow Fe^{2+}$ +0.77 V
$[Fe(CN)_6]^{3-} + e \rightarrow [Fe(CN)_6]^{4-}$ +0.36 V

From K_{stab}, $[Fe(CN)_6]^{3-}$ is more stable than $[Fe(CN)_6]^{4-}$, $[Fe(CN)_6]^{3-}$ is **less likely to be reduced** than Fe^{3+}(aq) or $[Fe(H_2O)_6]^{+3}$.
$[Fe(CN)_6]^{3-}$ is **stabilized by CN^-** ligands, resulting in a lower $E^\ominus$ value.

5 CJC/2009/P3/Q4(a)-(b)

4(a)

(i) $[H^+] = \sqrt{2.9 \times 10^{-14}} = 1.70 \times 10^{-7}$ mol dm^{-3}
$\therefore$ pH = - log (1.70×10^{-7}) = 6.77

(ii) Neutral, as $[H^+] = [OH^-]$

(iii) $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$

As temperature increases, K_w increases; so the equilibrium position shifts to the right. Therefore, the forward reaction (dissociation of water) is **endothermic**.

(b) $CH_3CH_2NH_3^+ + H_2O \rightleftharpoons CH_3CH_2NH_2 + H_3O^+$

$pK_b = 3.25$
$-\log K_b = 3.25$
$\therefore K_b = 5.62 \times 10^{-4}$ mol dm^{-3}
$\therefore K_a = \frac{K_w}{K_b} = \frac{2.9 \times 10^{-14}}{5.62 \times 10^{-14}} = 5.16 \times 10^{-11}$ mol dm^{-3}

$[H^+] = \sqrt{K_a \times c} = \sqrt{5.16 \times 10^{-11} \times 0.0200}$
$= 1.02 \times 10^{-6}$ mol dm^{-3}

pH = - log (1.02×10^{-6}) = 5.99

6 <u>DHS/2009/P2/Q4</u>

(a) Write an expression for the solubility product, K_{sp} of magnesium hydroxide.

$[Mg^{2+}][OH^-]^2$

(b) Calculate the solubility of magnesium hydroxide in g dm^{-3} for a saturated solution of magnesium hydroxide at 25°C.

Let the solubility be x.

$[Mg^{2+}][OH^-]^2 = 5.61 \times 10^{-12}$

$x(2x)^2 = 5.61 \times 10^{-12}$

$x = 1.119 \times 10^{-4}$ mol dm^{-3}

Solubility = 6.53×10^{-3} g

(c) Calculate the solubility of $Mg(OH)_2$ in 5.00×10^{-2} mol dm^{-3} aqueous sodium hydroxide.

Let the solubility of $Mg(OH)_2$ be x mol dm^{-3}.

$[Mg^{2+}][OH^-]^2 = 5.61 \times 10^{-12}$

$x(2x + 0.05)^2 = 5.61 \times 10^{-12}$

Assuming that x is so small such that $2x + 0.05 \sim 0.05$,

$x(0.05)^2 = 5.61 \times 10^{-12}$

$x = 2.24 \times 10^{-9}$ mol dm^{-3}

(d) Equal volumes of solutions containing 5.0×10^{-3} mol dm^{-3} magnesium nitrate and 6.0×10^{-3} mol dm^{-3} of sodium hydroxide are mixed. Predict if a precipitate will be formed. Explain your answer with the aid of relevant calculations.

$K_{sp} = [Mg^{2+}][OH^-]^2 = 5.61 \times 10^{-12}$ mol^2 dm^{-6}

On mixing aq $Mg(NO_3)_2$ and NaOH:

New $[Mg^{2+}]$ = V x 5.0 x 10^{-3} / 2V = 2.5 x 10^{-3} mol dm^{-3}
New $[OH^-]$ = V x 6.0 x 10^{-3} / 2V = 3 x 10^{-3} mol dm^{-3}

At saturation point, K_{sp} = Ionic product

Ionic product = $[Mg^{2+}][OH^-]^2$
= 9 x 10^{-8}[1] >5.61 x 10^{-12} [1]

Yes there will be a ppt.

7 **<u>DHS/2009/P3/1(a)</u>**

(i) Suggest a suitable indicator, if the titration were to be repeated without the use of a data logger. Explain your reasoning.

Phenolphthalein. The pH transition range of the indicator **lies within** the sharp pH change over its **equivalence point**.

(ii) Calculate the value of K_a for pyruvic acid.

No. of mol of NaOH = $\frac{30}{1000} \times 0.01$
= 3 x 10^{-4}
No. of mol of pyruvic acid = 3 x 10^{-4}
[pyruvic acid] = $\dfrac{3\times10^{-4}}{\frac{10}{1000}}$ = 0.03 mol dm^{-3}
From graph, pH = 2.5
Hence $[H^+] = 10^{-2.5}$ = 3.162 x 10^{-3}

$$K_a = \frac{[H^+]^2}{[pyruvic\_acid]} = \frac{(3.162\times10^{-3})^2}{0.03-3.162\times10^{-3}} = 3.73 \times 10^{-4} \text{ mol dm}^{-3}$$

(iii) Explain, with the aid of an appropriate equation, why the pH at equivalence point is greater than 7.

[7]

CH_3COCOO^- (aq) + H_2O (l) $\rightleftharpoons$ $CH_3COCOOH$ (aq) + OH^- (aq)
CH_3COCOO^- undergoes **salt hydrolysis**.
[OH] > [H$^+$].

8 HCI/2009/P2/Q1(b)

(b) (i) $[Ba^{2+}][F^-]^2 = 1.84 \times 10^{-7}$ where $[Ba^{2+}] = 0.05$ mol dm^{-3}

$\therefore [F^-] = \sqrt{(1.84 \times 10^{-7} / 0.05)}$

$= 1.92 \times 10^{-3}$ mol dm^{-3}

(ii) $[Ca^{2+}]_{remaining} = 3.45 \times 10^{-11} / (1.92 \times 10^{-3})^2$

$= 9.38 \times 10^{-6}$ mol dm^{-3}

9 HCI/2009/P2/Q2(a)

(a) (i) no. of moles of glycolic acid = 0.20/76.0 = 2.63×10^{-3} mol

volume of NaOH required = $2.63 \times 10^{-3}/0.10 \times 1000 = 26.3$ cm^3

(ii) no. of moles of CH_2OHCOO^- salt formed at equivalence = 2.63×10^{-3} mol

conc. of CH_2OHCOO^- = 2.63×10^{-3} /(26.3 + 20.0) = 0.0568 mol dm^{-3}

	CH_2OHCOO^-	+	H_2O	=	$CH_2OHCOOH$	+	OH^-
Initial conc. / mol dm^{-3}	0.0568				–		–
Eqm conc./ mol dm^{-3}	0.0568 - x				x		x

K_b of glycolate = K_w/K_a of glycolic acid = $1 \times 10^{-14} / 1.48 \times 10^{-4}$

$= 6.76 \times 10^{-11}$ mol dm^{-3}

$K_b = x^2/(0.0568 - x) \approx x^2/0.0568$ (assume $x \ll 0.0568$ mol dm^{-3})

$x = 1.959 \times 10^{-6}$ mol dm^{-3}

pOH = 5.708

pH = 14 – 5.708 = 8.29

(iii) Metacresol purple because its working range coincides with the sharp jump of the titration curve which lies in the alkaline pH region (equivalence pH = 8.29).

10 HCI/2009/P3/Q5(a)

5 (a) (i) A solution that is able to resist pH changes when small quantities of acid or base are added.

(ii) When a small amount of OH^- is added, OH^- ions can be removed from the system by the reaction: $CO_2 + OH^- \rightarrow HCO_3^-$

{or $H_2CO_3 + OH^- \rightarrow HCO_3^- + H_2O$}

Hence pH remains almost constant.

(iii) $7.90 \times 10^{-7} = 10^{-7.4}\ [HCO_3^-]/[CO_2]$

$[HCO_3^-]/[CO_2]$ is 20:1

Blood has greater capacity for absorbing H^+ since there is a much higher concentration of the base component HCO_3^-.

(iv) CO_2 exhaled is recycled. Blood CO_2 level is raised and the buffer equilibrium shifts to the right thus increasing $[H^+]$. Hence blood pH will drop.

11 IJC/2009/P2/Q4(a)

(i) Explain what is meant by the term base dissociation constant, K_b of morphine.

$$Mor + H_2O \rightleftharpoons MorH^+ + OH^-$$

$$K_b = \frac{[MorH^+][OH^-]}{[Mor]}$$

(ii) Calculate the pH of 0.20 mol dm^{-3} morphine solution.

At equilibrium, $[OH^-] = [MorH^+] = x$ mol dm^{-3}

$[Mor] = 0.20$ mol dm^{-3} (assume $x << 0.20$)

Now, $7.40 \times 10^{-7} = \frac{x^2}{0.20}$

$x = 3.85 \times 10^{-4}$ mol dm^{-3}

i.e. $pOH = -\lg(3.85 \times 10^{-4}) = 3.41$

$pH = 14 - 3.41 = \underline{\underline{10.6}}$

(iii) Calculate the amount of salt formed in the buffer solution.

$$Mor + H^+ \longrightarrow MorH^+$$

$n_{MorH^+} = n_{HCl} = \underline{\underline{0.01 \text{ mol}}}$

(iv) Calculate the pH of the buffer solution.

$n_{Mor} = \frac{100}{1000} \times 0.2 - 0.01 = 0.01$ mol

$\therefore [Mor] = [MorH^+] \Rightarrow pOH = pK_b$

$= 6.13$

$\therefore pH = 14 - 6.13$

$= \underline{\underline{7.87}}$

12 JJC/2009/P3/Q2(a)-(b)

2. (a) (i) A weak acid is a **proton donor** that **dissociates partially**.

(ii) $[H^+]$ = **5.66×10^{-3}** mol dm^{-3} **[1]**

pH = **2.25** **[1]**

(iii) Amount of acid used = 5×1^{-3} mol

Amount of NaOH required = 5×10^{-3} mol

Volume of NaOH required = **$0.0500\ dm^3 = 50\ cm^3$**

$[OH^-] = \sqrt{K_b \text{ of acetylsalicylic acid} \times \text{conc of salt solution}}$

$= \sqrt{\dfrac{1.00\times10^{-14}}{3.2\times10^{-4}} \times \dfrac{5\times10^{-3}}{0.05+0.05}}$ **[1]**

$= 1.25 \times 10^{-6}$ mol dm^{-3}

pOH = 5.90

pH = **8.10** **[1]**

(b) (i) pH = pK_a = $-\lg(3.2\times10^{-4})$ = **3.49** **[1]**

(ii) When H^+ is added:

$C_8H_7O_2COO^- + H^+ \rightarrow C_8H_7O_2COOH$ **[1]**

When OH^- is added:

$C_8H_7O_2COOH + OH^- \rightarrow C_8H_7O_2COO^- + H_2O$ **[1]**

(iii) Amount of HC*l* added = 2.00×10^{-4} mol

In the resultant solution:

Amount of $C_8H_7O_2COO^-$ left = $2.5\times10^{-3} - 2\times10^{-4} = 2.30\times10^{-3}$ mol

Concentration of $C_8H_7O_2COO^-$ = $\dfrac{2.30\times10^{-3}}{50+25+2 / 1000}$

= 0.0299 mol dm^{-3}

2. (b) (iii) Amount of $C_8H_7O_2COOH$ = $\underline{2.5\times10^{-3}+2\times10^{-4}}=2.70\times10^{-3}$ mol

Concentration of $C_8H_7O_2COOH$ = $\dfrac{2.70\times10^{-3}}{50+25+2/1000}$

= 0.0351 mol dm^{-3}

pH of the resulting solution = $-\lg\left(3.2\times10^{-4}\right)+\lg\dfrac{0.0299}{0.0351}$

= **$\underline{3.43}$**

13 <u>MI/2009/P2/Q2(b)</u>

(i) Calculate the amount of H_2CO_3 present in the sample of blood plasma.

At the first endpoint, ***$H_2CO_3(aq) + NaOH(aq) \rightarrow NaHCO_3(aq) + H_2O(l)$***

Amount of H_2CO_3 present = Amount of NaOH used for the 1st endpoint

$=\dfrac{1.30}{1000}\times7.50\times10^{-3}$

$=9.75\times10^{-6}$ mol

(ii) Calculate the amount of HCO_3^- present in the sample of blood plasma. [2]

At the second endpoint, ***$NaHCO_3(aq) + NaOH(aq) \rightarrow Na_2CO_3(aq) + H_2O(l)$***

Note: Amount of HCO_3^- present for step 2 of the titration

= Amount of HCO_3^- present originally + Amount of HCO_3^- produced from step 1 of the titration

= Amount of HCO_3^- present originally + 9.75 x 10^{-6}

Amount of HCO_3^- present altogether

= Amount of NaOH used for the 2nd endpoint only

$=\dfrac{14.60}{1000}\times7.50\times10^{-3}$;

$=1.095\times10^{-4}$ mol

Amount of HCO_3^- present originally $=1.095\times10^{-4}-9.75\times10^{-6}$;

$=9.98\times10^{-5}$ mol

(iii) Hence determine the K_a of carbonic acid.

$$\text{pH of buffer} = pK_a + \lg\frac{[HCO_3^-]}{[H_2CO_3]}$$

$$7.4 = pK_a + \lg\frac{\left(9.975\times10^{-5}/10.0\right)}{\left(9.75\times10^{-6}/10.0\right)}$$

$$pK_a = 7.4 - \lg\left(\frac{9.975\times10^{-5}}{9.75\times10^{-6}}\right)$$

$$pK_a = 6.39$$

$$K_a = 4.07\times10^{-7}\ \text{mol dm}^{-3}$$

(iv) The theoretical value for the K_a of carbonic acid is 4.47 x 10^{-7} mol dm^{-3}. Suggest an error or limitation of this experiment which would have resulted in the difference in the value calculated in (b)(iii). [1]

There is a large percentage error in the titration values recorded as the volume of NaOH used for the first endpoint is very small. ;

OR

$$\%\ \text{error} = 2\times\frac{\frac{1}{2}\times0.10}{1.30}\times100\% = 7.69\%$$

(v) Using the theoretical value for the K_a of carbonic acid, calculate the pH of the first endpoint. [4]

$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$$

$$K_w = K_a\times K_b$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0\times10^{-14}}{4.47\times10^{-7}} = 2.237\times10^{-8}\ \text{mol dm}^{-3}$$

$$[HCO_3^-] = \frac{1.095\times10^{-4}}{10.0}\times1000 = 0.01095\ \text{mol dm}^{-3}$$;;

$$K_b = \frac{[H_2CO_3][OH^-]}{[HCO_3^-]} = \frac{[OH^-]^2}{[HCO_3^-]}$$

$$[OH^-] = \sqrt{2.237\times10^{-8}\times0.01095} = 1.565\times10^{-5}\ \text{mol dm}^{-3}$$

$$\text{pH} = 14 - \text{pOH} = 14 - \left[-\lg(1.565\times10^{-5})\right] = 9.19$$;;

14 MJC/2009/P2/Q2

(a) Determine whether a precipitate is formed when the common washing-up liquid was accidentally mixed with 0.200 dm^3 of 'hard' water.

Ionic product of $(C_{18}H_{19}SO_3)_2Ca = (7.143 \times 10^{-5})(3.643 \times 10^{-5})^2$

$= \underline{\mathbf{9.48 \times 10^{-14}\ mol^3\ dm^{-9}}}$

Ionic product > K_{sp} $\Rightarrow$ precipitate of $(C_{18}H_{19}SO_3)_2Ca$ is formed.

(i) Write the K_c expression for the reaction.

$$K_c = \frac{[CaP_3O_{10}^{3-}]}{[Ca^{2+}][P_3O_{10}^{3-}]}$$

(ii) Hence, calculate the concentration of tripolyphosphate ions required to reduce the calcium ion concentration in a typical sample of "hard water" to 1.0×10^{-6} mol dm^{-3}.

Let x be the initial concentration of $P_3O_{10}^{5-}$ required

At equilibrium,

$$7.7 \times 10^8 = \frac{2.49 \times 10^{-4}}{(1.0 x 10^{-6})(x - 2.49 \times 10^{-4})}$$

$\underline{\mathbf{x = 2.493 \times 10^{-4}\ mol\ dm^{-3}}}$

15 MJC/2009/P3/Q4(a)

4(ai) Ka_1 and Ka_2 are the acid dissociation constant of the carboxylic acid group and phenolic group respectively. Ka_1 is smaller Ka_2 because carboxylic acid group is more acidic than the phenolic group.

(ii) H^+ from the 2nd dissociation is negligible and can be ignored and all the H^+ comes from the 1st dissociation.

$[H^+] = \sqrt{cK_a}$

$= \sqrt{0.05 \times 6.31 \text{ x } 10^{-5}}$

$= 1.78 \text{ x } 10^{-3}$ mol dm^{-3}

Hence, pH = **2.75**

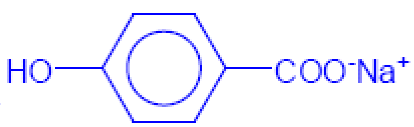
(iii) For acidic buffer

$$pH = pKa + \lg\frac{[salt]}{[acid]}$$

$$= -\lg 6.31 \text{ x } 10^{-5} + \lg\frac{[6.25\times10^{-4}/0.03125]}{[6.25\times10^{-4}/0.03125]}$$

= **4.20**

Salt hydrolysis

[

] = $1.25 \text{ x}10^{-3} / 0.0375 = 3.33 \text{ x } 10^{-2}$ mol dm^{-3}

$$K_{b1} = \frac{K_w}{K_{a1}} = \frac{1\times10^{-14}}{6.31\times10^{-5}} = 1.58\times10^{-10} \text{ mol dm}^{-3}$$

$[OH^-] = \sqrt{cK_b}$

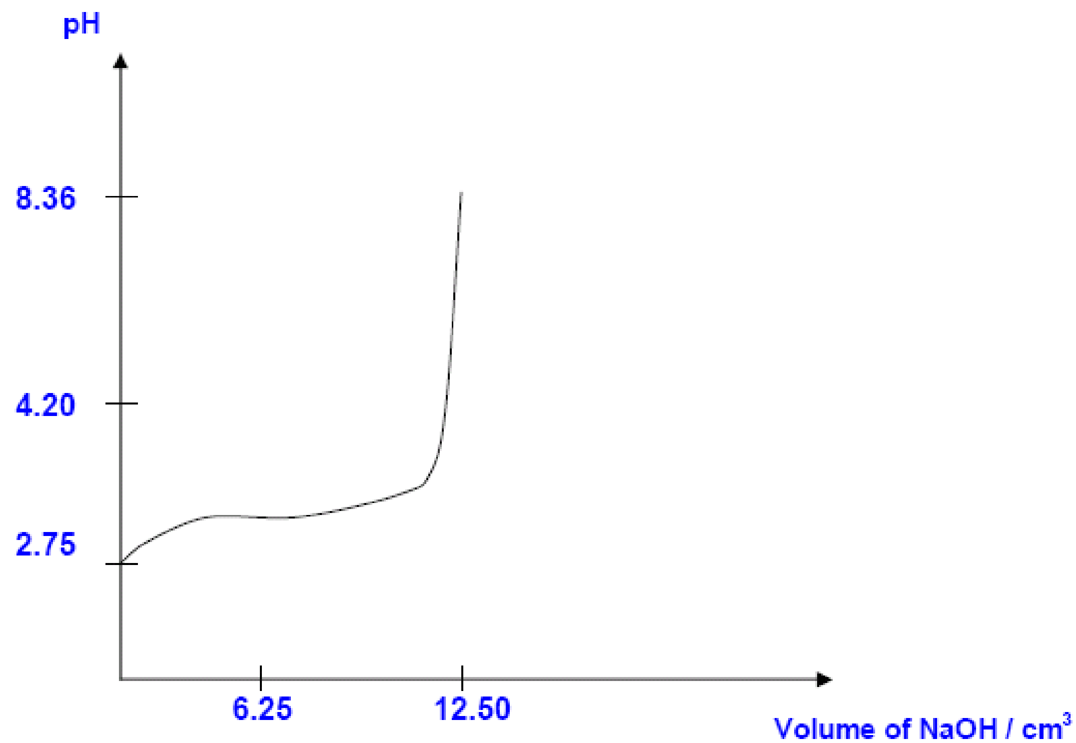
$= \sqrt{3.33\times10^{-2}\times1.58\times10^{-10}}$

$= 2.29\times10^{-6}$ mol dm^{-3}

pOH = $-\lg 2.29 \text{ x } 10^{-5} = 5.63$

pH = 14 – 5.63 = **8.36**

(iv)



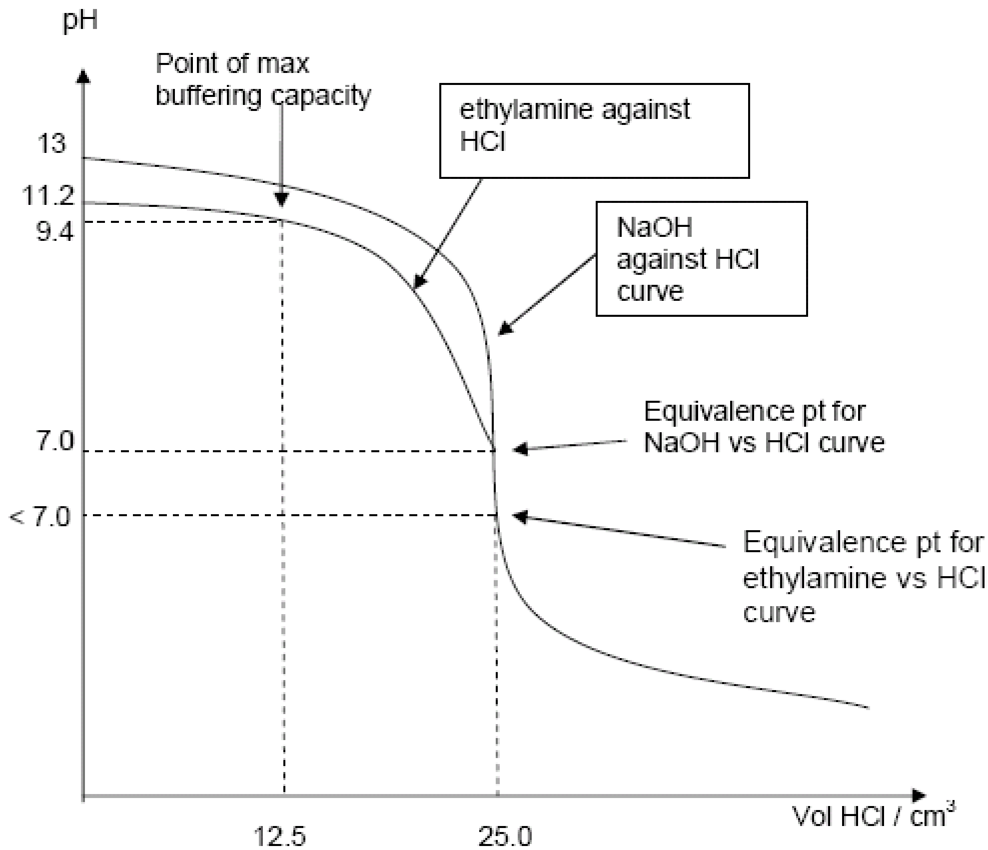
16 **NJC/2009/P3/Q2(a)**

2 (i)(I) Minimum $[Ag^+] = 8.52\times10^{-14}\,moldm^{-3}$

(II) Minimum $[Ag^+] = 3.17\times10^{-5}\,moldm^{-3}$

(ii) $[I^-] = 2.69\times10^{-12}\,moldm^{-3}$

17 NJC/2009/P3/Q5(e)(ii)



18 NYJC/2009/P3/Q4(a)-(b)

4(a) **(i)** $[OH^-] = \sqrt{5.6 \times 10^{-4}(0.200)} = 1.058 \times 10^{-2}$ mol dm^{-3}

pOH = - lg 1.058 x 10^{-2} = 1.98 pH = 12.02

(ii) Using $\dfrac{0.200 \times V_{CH_3CH_2NH_2}}{0.120 \times 40.0} = \dfrac{1}{1} \Rightarrow V_{CH_3CH_2NH_2} = 24.0$ cm^3

(iii) When vol = 48.0 cm^3 , solution reached max buffering capacity

(b) **(i)** $PbI_2 \rightarrow Pb^{2+} + 2I^-$

Let **x** be $[I^-]$

$9.8 \times 10^{-9} = \frac{1}{2}x(x)^2$

$x = \sqrt[3]{2(9.8 \times 10^{-9})} = 2.696 \times 10^{-3}$ mol dm^{-3}

Solubility of PbI_2 = $\dfrac{2.696 \times 10^{-3}}{2}$ mol dm^{-3}

(ii) Let C be $[Pb^{2+}]$ before mixing

$[Pb^{2+}]$ after mixing $= \frac{C}{2}$ $[I^-]$ after mixing $= \frac{0.500}{2}$ [1m]

for ppt to appear, IP of PbI_2 $>$ K_{sp} of PbI_2

$$\left(\frac{C}{2}\right)\left(\frac{0.500}{2}\right)^2 > 9.8 \times 10^{-9}$$ [1m]

$\Rightarrow C > 3.136 \times 10^{-7}$ mol dm^{-3}

19 PJC/2009/P2/Q3(b)

(b) (i) $Fe(IO_3)_3(s) + aq \rightleftharpoons Fe^{3+}(aq) + 3IO_3^-(aq)$

$[Fe^{3+}] = 3.6/(580.8) = 6.20 \times 10^{-3}$ mol dm^{-3}
$[IO_3^-] = 3 \times 6.20 \times 10^{-3} = 0.0186$ mol dm^{-3}

K_{sp} of $Fe(IO_3)_3 = [Fe^{3+}]_{eqm} \times [IO_3^-]^3_{eqm}$
$= 3.99 \times 10^{-8}$ mol^4 dm^{-12} at 298 K

(ii) Due to the Common Ion Effect, the solubility decreases.

Let the solubility be x mol dm^{-3}
$[IO_3^-]_{total} = 0.105 + 3x \approx 0.105$ mol dm^{-3}; since 3x is small compared to 0.105

K_{sp} of $Ce(IO_3)_3 = [Ce^{3+}]_{eqm} \times [IO_3^-]^3_{eqm}$
3.99×10^{-8} mol^4 dm^{-12} $= x(0.105)^3$ mol^4 dm^{-12}
$x = 3.44 \times 10^{-5}$ (check: $3x = 1.03 \times 10^{-4} << 0.105$)

Solubility of $Fe(IO_3)_3$ in 0.100 mol dm^{-3} KIO_3 = **3.44 x 10^{-5} mol dm^{-3}**.

20 PJC/2009/P3/Q1(a)

1(a)(i) $FeBr_3(s) + 6H_2O \rightarrow [Fe(H_2O)_6]^{3+} + 3Br^-$

$[Fe(H_2O)_6]^{3+} + H_2O \rightleftharpoons [Fe(H_2O)_5OH]^{2+} + H_3O^+$

Due to high charge density of Fe^{3+} ion, it is able to polarize the water molecule weakening the O—H bond, releasing H^+ ion. Hence a solution of $FeBr_3$ has pH less than 7 at 298 K.

(ii) K_a (Fe^{3+} (aq)) $= 6.34 \times 10^{-6}$ mol dm^{-3}

21 PJC/2009/P3/Q3(a)

3(a) (i) $K_a = 10^{-pKa} = 10^{-6.80} = 1.58 \times 10^{-7}$ moldm^{-3}

$K_a = [C_{10}H_{14}N]\,[H_3O^+] / [C_{10}H_{14}NH^+]$

$K_a = [H_3O^+]^2 / [C_{10}H_{14}NH^+]$

$[H_3O^+] = [(1.58 \times 10^{-7})(0.100)]^{1/2} = 1.26 \times 10^{-4}$ moldm^{-3}

pH = $-\log_{10}[H_3O^+] = 3.90$

(ii) When pH is decreased, $[H_3O^+]$ increases and position of equilibrium shift to the left, concentration of conjugate of nicotine increases.

(iii) pH = $pK_a(C_{10}H_{14}NH^+) + \log_{10} [C_{10}H_{14}N]/[C_{10}H_{14}NH^+] = 6.50$

(iv) $[C_{10}H_{14}N]_{new} = (5 \times 0.100 - 0.200v) / (5.00 + v)$

$[C_{10}H_{14}NH^+]_{new} = (5 \times 0.200 + 0.200v) / (5.00 + v)$

pH = $pK_a(C_{10}H_{14}NH^+) + \log_{10} [C_{10}H_{14}N]/[C_{10}H_{14}NH^+]$

$$6.40 = 6.80 + \log_{10}\frac{(5 \times 0.100 - 0.200v)/(5.00 + v)}{(5 \times 0.200 + 0.200v)/(5.00 + v)}$$

solving, v = **0.364 dm^3**

22 RI/2009/P2/Q1(c)

(c)(i) $2H^+(aq) + Mg(OH)_2(s) \longrightarrow Mg^{2+}(aq) + 2H_2O(l)$

(c)(ii) $\mathbf{Mg^{2+}(aq) + 2F^-(aq) \longrightarrow MgF_2(s)}$

The Mg^{2+} released from the acid–base reaction in the stomach reacts with the F^- present to form a precipitate of MgF_2 which is **insoluble** and is thus **not easily absorbed** by the body.

(c)(iii) Precipitation occurs when

$[Mg^{2+}][F^-]^2 > K_{sp}$

$[Mg^{2+}][(\frac{1\times10^{-3}}{19.0})/1.0]^2 > 5.16 \times 10^{-11}$

$[Mg^{2+}] > 0.0186$ mol dm^{-3}

(vol. of milk of magnesia needed)(1.40) = (0.0186)(vol. of liquid in stomach)

i.e. vol. of milk of magnesia needed = 0.0133 dm^3

(i.e. 13.3 cm^3)

23 <u>RI/2009/P3/Q1(a)</u>

(a)(i) Amount of lactic acid = Amount of NaOH needed for neutralisation

$$= \frac{20.00}{1000} \times 0.120 = 0.0024 \text{ mol}$$

Concentration of lactic acid = $0.0024 \div \frac{25.0}{1000} = 0.0960 \text{ mol dm}^{-3}$

(ii) pK_a = pH at half-neutralisation = 3.85 (when V_{NaOH} = 10 cm^3)
$K_a = 10^{-3.85} = 1.41 \times 10^{-4}$ mol dm^{-3}

Alternatively, $K_a = \frac{[H^+][A^-]}{[HA]}$

Initially, $[H^+] = [A^-] = 10^{-2.45} = 3.548 \times 10^{-3}$ mol dm^{-3}

$K_a = \frac{3.548 \times 10^{-3} \times 3.548 \times 10^{-3}}{0.0960 - 3.548 \times 10^{-3}} = 1.36 \times 10^{-4}$ mol dm^{-3} where HA = lactic acid

(iii) At equivalence point, the resultant mixture is a solution of sodium lactate, $H_3CH(OH)COO^-Na^+$.
Being the conjugate base of a weak acid, lactate ion hydrolyses in water to give OH^-:

$CH_3CH(OH)COO^-(aq) + H_2O(l) \rightleftharpoons CH_3CH(OH)COOH(aq) + OH^-(aq)$

The formation of OH^- causes the pH at equivalence point to be greater than 7.

24 <u>SAJC/2009/P2/Q6</u>

$$Ca(OH)_2 (s) \rightleftharpoons Ca^{2+} (aq) + 2OH^- (aq)$$
- s +s +2s

(a) $Ksp = s(2s)^2 = 4s^3 = 1.0 \times 10^{-12}$
s= 6.30×10^{-5}, solubility in water is 6.30×10^{-5} mol dm^{-3}.

(b) $[Ca^{2+}] = (0.10 + s)$
$[OH^-] = 2s$
$Ksp = (0.10 + s)(2s)^2$
$1.1 \times 10^{-12} = (0.10 + s)(2s)^2$
$s = 1.58 \times 10^{-6}$ mol dm^{-3}

(c) $[Ca^{2+}] = 1\times10^{-3}/2 = 0.5 \times 10^{-3}$ mol dm^{-3}
$[OH^-] = 2 \times 2 \times 10^{-3} /2 = 2 \times 10^{-3}$ mol dm^{-3}
Ionic product = $0.5 \times 10^{-3} \times (2 \times 10^{-3})^2 = 2 \times 10^{-9}$ mol^3 dm^{-9} > Ksp
Hence, precipitation takes place.

25 SAJC/2009/P3/Q4

4. (a) (i) pH of lactic acid = 2.5

$[H^+] = 10^{-2.}$

Since lactic acid is a weak monobasic acid,

$K_a = [H^+]^2 / [CH_3CHOH(COOH)]$ OR $[H^+] = \sqrt{K_a[HA]}$

$= (10^{-2.5})^2 / 0.080$

$= 1.28 \times 10^{-4}$ mol dm^{-3}

(ii) Maximum buffer capacity occurs when [salt]=[acid]

$pH = pK_a$

$= -\log 1.25 \times 10^{-4}$

$=3.90$

(iii) When a small amount of H^+ is added,

$CH_3CH(OH)COO^-\ (aq) + H^+\ (aq) \rightleftharpoons CH_3CH(OH)COOH$

The additional acid, H^+, is removed by <u>large concentration of $CH_3CH(OH)COO^-$</u> from the <u>salt</u>.

Thus, H^+ changes very slightly and the <u>pH remains almost constant.</u>

(iv) At the equivalence point, only basic salt is present.

No. of moles of salt formed = $0.08 \times 10/1000 = 8 \times 10^{-4}$ mol

[salt] = $8 \times 10^{-4} / 26 \times 1000 = 0.031$ mol dm^{-3}

$[OH^-] = \sqrt{[(1 \times 10^{-14} / 1.25 \times 10^{-4}) \times 0.031]} = 1.55 \times 10^{-6}$ mol dm^{-3}

pOH = 5.8

pH = 8.2

A suitable indicator is <u>phenolphthalein</u>.

26 SRJC/2009/P2/Q4

(a) Explain the term acid dissociation constant, K_a, as applied to ethanoic acid. **[1]**

Acid dissociation constant, K_a is a measure of the **strength of a weak acid**
OR
$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$ **[1M]**

(b) Calculate the initial pH of the 20.0 cm^3 sample of ethanoic acid. **[2]**

$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$
Let the $[H^+]$ be x mol dm^{-3}

$$\frac{(x)(x)}{0.01 - x} = K_a$$

Since $K_a < 10^{-4}$ mol dm^{-3}; assume that x is so small that $0.01 - x \approx 0.01$

$$\frac{(x)(x)}{0.01} = 1.8\times10^{-5}$$ **[1 M]**

$[H^+]$ = **4.242×10^{-4} mol dm^{-3}** **[1/2 M]**
pH = $-\lg[4.24 \times 10^{-4}]$
= **3.37** **[1/2 M]**

(c) Calculate the equivalence volume of NaOH and hence, the end point pH. **[4]**

$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COO^-Na^+(aq) + H_2O(l)$

Amount of CH_3COOH = $\frac{20}{1000}\times 0.01$
= 2.000×10^{-4} mol

Since $n_{CH3COOH} : n_{NaOH}$ = 1:1

Amount of NaOH = **2.000×10^{-4} mol** **[1/2 M]**

Equivalence volume of NaOH = $\frac{2\times10^{-4}}{0.02}$
= **0.0100 dm^3** = **10.0 cm^3** **[1M]**

At equivalence point,

$[CH_3COO^-]$ = $\frac{2\times10^{-4}}{30/1000}$
= 6.667×10^{-3} mol dm^{-3} **[1/2 M]**

Let the $[OH^-]$ at equivalence point be y mol dm^{-3}

	$CH_3COO^-(aq)$ +	$H_2O(l)$ $\rightleftharpoons$	$CH_3COOH(aq)$ +	$OH^-(aq)$
Initial [ ]:	6.667×10^{-3}	-	0	0
Δ in [ ]:	- y	-	+ y	+ y
Eqm [ ]:	$6.667 \times 10^{-3} - y$	-	y	y

$K_a \times K_b = K_w$

$K_b = \dfrac{10^{-14}}{1.8\times10^{-5}}$

$= 5.556 \times 10^{-10}$ mol dm^{-3} **[1/2 M]**

$$\frac{(y)(y)}{(6.667\times10^{3})-y} = K_b$$

Since $K_b < 10^{-4}$ mol dm^{-3}; assume that x is so small that $6.667 \times 10^{-3} - x \approx 6.667 \times 10^{-3}$

$$\frac{(y)(y)}{6.667\times10^{3}} = 5.556\times10^{-10}$$ **[1/2 M]**

$[OH^-]$ = 1.924×10^{-6} mol dm^{-3}

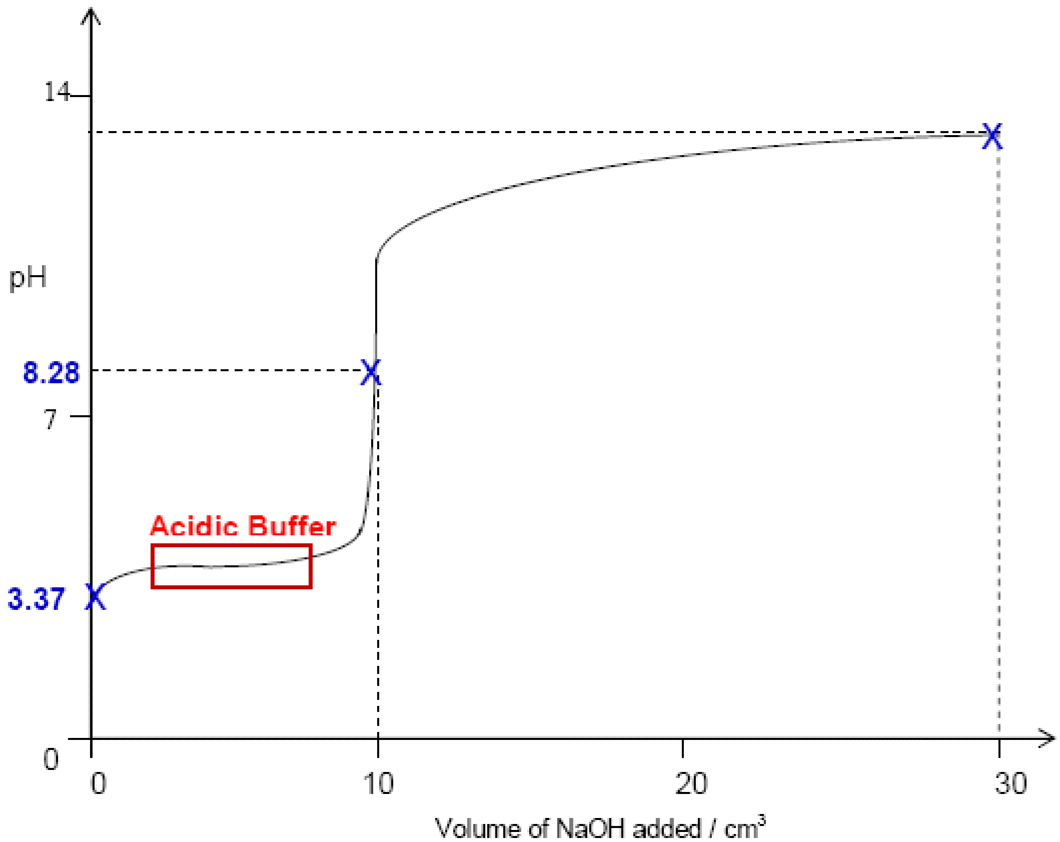
pOH = -lg (1.924×10^{-3}) = 5.716 **[1/2 M]**

pH = 14 – 5.716

= **8.28** **[1/2 M]**

[4]

(d)	The reaction is continued until 30 cm^3 of sodium hydroxide has been added. On the given grid, using the pH values you have determined from **(b)** and **(c)**, sketch how the pH changes and indicate clearly the buffer region. **[3]**



Mark all 3 points (initial, endpoint and final pH) correctly – **[1M]**
Correct shape of the graph – **[1M]**
Correct indication of buffer region – **1M]**

(e)	State and explain the choice of a suitable indicator for this reaction. **[2]**

<u>Phenolphthalein.</u> **[1M]**

Its **<u>pH transition range</u>** (≈ 8 – 9.8) **<u>lies within</u>** the sharp **<u>pH change</u>** over the **<u>equivalence point</u>**. **[1M]**

27 SRJC/2009/P3/Q3(c)

(i) A sample of sodium hypochlorite, NaOC*l*, was dissolved in 100 cm^3 of 0.123 mol dm^{-3} HOC*l* solution forming a buffer of pH 6.20. Determine the ratio of the concentration of OCl^- to HOC*l* in the solution. [1]

$$pH = pKa + lg\frac{[OCl^-]}{[HOCl]}$$

$$6.20 = 7.50 + lg\frac{[OCl^-]}{[HOCl]}$$

$$lg\frac{[OCl^-]}{[HOCl]} = -1.30$$

$$\frac{[OCl^-]}{[HOCl]} = 0.0501$$

(ii) The buffer is then used to absorb HC*l* gas. By using the above ratio or otherwise, calculate the amount of gaseous HC*l* (in mol) that is required to be added to the buffer solution until it reaches pH 6. [4]

$$\frac{[OCl^-]}{[HOCl]} = 0.0501$$

$$[OCl^-] = 0.0501[HOCl]$$

$$= 0.0501 \times 0.123$$

$$= 0.006162 \text{ mol dm}^{-3}$$

On addition of HCl, let x be the concentration of H^+

$[HOCl]_{new} = 0.123 + x$

$[OCl^-]_{new} = 0.006162 - x$

$$6.00 = 7.50 + lg\frac{[OCl^-]}{[HOCl]}$$

$$6.00 = 7.50 + lg\frac{0.006162 - x}{0.123 + x}$$

$$-1.50 = lg\frac{0.006162 - x}{0.123 + x}$$

$$0.03162 = \frac{0.006162 - x}{0.123 + x}$$

$0.03162(0.123) + 0.03162x = 0.006162 - x$

$1.03162x = 0.00227$

$x = 0.00220 \text{ mol dm}^{-3}$

$$\text{amount of HCl} = 0.00220 \times \frac{100}{1000}$$

$$= 0.000220 \text{ mol}$$

28 TJC/2009/P3/Q1(d)

(d) (i)

- $K_{sp} = [Pb^{2+}][CrO_4^{2-}]$

  Let the solubility of $PbCrO_4$ be x mol dm^{-3}

  $$PbCrO_4(s) \rightleftharpoons Pb^{2+}(aq) + CrO_4^{2-}(aq)$$

- $K_{sp} = [Pb^{2+}][CrO_4^{2-}]$

  $1.69 \times 10^{-14} = x^2$

  $x = \sqrt{K_{sp}}$

- $= 1.30 \times 10^{-7}$ **mol dm^{-3}**

The solubility of $PbCrO_4$ is 1.30×10^{-7} mol dm^{-3}.

(ii)

- $[Pb^{2+}] = \dfrac{K_{sp}}{[CrO_4^{2-}]} = \dfrac{1.69\times10^{-14}}{0.010}$
- $= 1.69 \times 10^{-12}$ mol dm^{-3}

29 TJC/2009/P3/Q3

(a) (i)

- At the surface of the sea, pressure of CO_2 is 1 atm (normal atmospheric pressure)

  At 300 m below sea level, the pressure of CO_2 = $\dfrac{300}{10}+1$

  = 31 atm

- Solubility of CO_2 300 m below sea level

  $= \dfrac{31}{1}\times3.29\times10^{-2}$

  = 1.02 mol dm^{-3}

(ii) (I)

- $[H^+] = \sqrt{K_a \times C}$

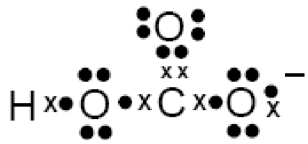
  $= \sqrt{4.5\times10^{-7}\times1.02}$

  $= 6.77 \times 10^{-4}$ mol dm^{-3}

- pH = $-\lg(6.77 \times 10^{-4})$

  = 3.17

(II)

- 

(III)
- Trigonal planar

(iii)
(I)
- $K_{a2} = \dfrac{[HCO_3^-][H^+]}{[H_2CO_3]}$

(II)

$$K_{a1} = \frac{[HCO_3^-][H^+]}{[CO_2]}$$
$$= 4.5 \times 10^{-7} \text{ mol dm}^{-3}$$

$$K_{a2} = \frac{[HCO_3^-][H^+]}{[H_2CO_3]}$$

- $= \dfrac{[HCO_3^-][H^+]}{[CO_2]} \times \dfrac{[CO_2]}{[H_2CO_3]}$

$= 4.5 \times 10^{-7} \times 400$

- $= 1.80 \times 10^{-4} \text{ mol dm}^{-3}$

30 TPJC/2009/P3/Q3(a)-(b)

3a)i) $B + H_2O \rightleftharpoons BH^+ + OH^-$

ii) In the neutralization rx:

HCl +	B →	BH^+	Cl^-

At the end of rx: there is salt, BH^+Cl^- , water and excess B to make a buffer soln.
Let the vol HCl = x dm^3
Vol of B used = (1-x) dm^3

	HCl +	B →	BH^+	Cl^-
Initial /mol	0.05x	0.05(1-x)	-	-
Change	-0.05x	-0.05x	+0.05x	+0.05x
Final/ mol	-	0.05(1-x)- 0.05x = 0.05-0.1x	0.05x	0.05x

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

$[OH^-] = 10^{-14}/10^{-9.2} = 10^{-4.8}$
$K_b = 3.17 \times 10^{-5} = \frac{[BH^+]}{[B]} \times 10^{-4.8}$
$2.00 = 0.05x/0.05-0.1x$
$0.1 - 0.2x = 0.05x$
x = 0.400 dm^3 , vol HCl;
vol B = 0.600 dm^3

iii) Dilution has no effect on the ration of $\frac{[BH^+]}{[B]}$ and K_b, hence pH remains the same at 9.2.

3bi) Eqm 1: $AgCl(s) + 2NH_3(aq) \rightarrow [Ag(NH_3)_2]^+ Cl^-(aq)$
Eqm 2: $AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$

refer eqm1: Addition of ammonia soln form the complex, $[Ag(NH_3)_4]^+$, hence reducing the $[Ag^+(aq)]$,
this affect the eqm 2, by LCP, system tries to replenish the $Ag^+(aq)$ by causing eqm 2 to shift to the right, resulting in more AgCl to dissolve.

ii) In 2 dm^3 soln,
Amt of AgCl dissolved
$= 5.00 \times 10^{-5} - 1.88 \times 10^{-5} = 3.12 \times 10^{-5}$ mol
Let amt of complex Ag^+ be y mol.

$Ksp = [Ag^+][Cl^-]$
$1.8 \times 10^{-10} = \frac{(3.12 \times 10^{-5} - y)}{2.00} \times \frac{(3.12 \times 10^{-5})}{2.00}$
$y = 0.810 \times 10^{-5}$ mol in 2.00 dm^3

hence, [complexed Ag^+] $= 0.810 \times 10^{-5} / 2.00$
$= \underline{4.05 \times 10^{-6}}$ mol dm^{-3}

31 VJC/2009/P2/Q2(b)

(i) Calculate the concentration of hydrogen ions in the solution after barium hydroxide is added.

$\mathbf{[H^+] = 10^{-2.88} = 1.32 \times 10^{-3}}$ **mol dm^{-3}**

(ii) Using your answer to **(i)** and the given information, calculate the value of the acid dissociation constant, K_a, of HF.

$\mathbf{K_a = [H^+][F^-] / [HF]}$
$\mathbf{= 1.32 \times 10^{-3} \times (1/2)}$
$\mathbf{= 6.60 \times 10^{-4}}$ **mol dm^{-3}**

32 <u>VJC/2009/P3/Q2(a)</u>

(i) Write an expression for the solubility product of $Ca(OH)_2$.

$$Ca(OH)_2\,(s) \rightleftharpoons Ca^{2+}\,(aq) + 2OH^-\,(aq)$$

K_{sp} of $Ca(OH)_2$ = $[Ca^{2+}][OH^-]^2$

(ii) Calculate the pH of a saturated solution of $Ca(OH)_2$ at 25°C given that its solubility is 0.830 g dm^{-3}.

M_r of $Ca(OH)_2$ = 74.1

Solubility of $Ca(OH)_2$ = $\frac{0.830}{74.1}$ = 1.12×10^{-2} mol dm^{-3}

$[OH^-] = 2 \times 1.12 \times 10^{-2} = 2.24 \times 10^{-2}$ mol dm^{-3}

pH = 14 – lg(2.24×10^{-2}) = 12.35

(ii) Determine the solubility product of $Ca(OH)_2$ at 25°C, stating its units.

$$K_{sp} \text{ of } Ca(OH)_2 = [Ca^{2+}][OH^-]^2 = (1.12 \times 10^{-2})(2.24 \times 10^{-2})^2 = 5.62 \times 10^{-6}\ mol^3\,dm^{-9}$$

(iii) Explain and predict qualitatively the effect(if any) on the solubility and solubility product of $Ca(OH)_2$ when 25.0 cm^3 of 0.050 mol dm^{-3} solution of potassium hydroxide is added to the solution in (ii)

The <u>common ion, OH^-,</u> from the very soluble KOH would cause <u>eqm in (i) to shift left</u>, hence <u>lowering the solubility</u> of $Ca(OH)_2$.
K_{sp} of $Ca(OH)_2$ would <u>not be affected</u> as <u>temperature is constant</u>.

(iv) Write a balanced equation to show the reaction between carbon dioxide and baryta water.

$$Ba(OH)_2 + CO_2 \longrightarrow BaCO_3 + H_2O$$

(v) Which solution, lime or baryta water,is more sensitive to the detection of carbon dioxide gas at 25°C? Explain your answer with reference to the data given in table in (a).

As the <u>K_{sp} of $CaCO_3$ is lower</u>, the <u>ionic product</u> of $[Ca^{2+}][CO_3^{2-}]$ would <u>exceed the K_{sp} of $CaCO_3$ more readily</u> hence <u>lime water would be more sensitive</u>.

[2]

33 YJC/2009/P2/Q3(c)

(c) **(i)** $[H^+] = \sqrt{K_a \times c} = \sqrt{2.05 \times 10^{-5} \times 0.100} = 1.43 \times 10^{-3}\, moldm^{-3}$

$pH = -\log(1.43 \times 10^{-3}) = 2.84$

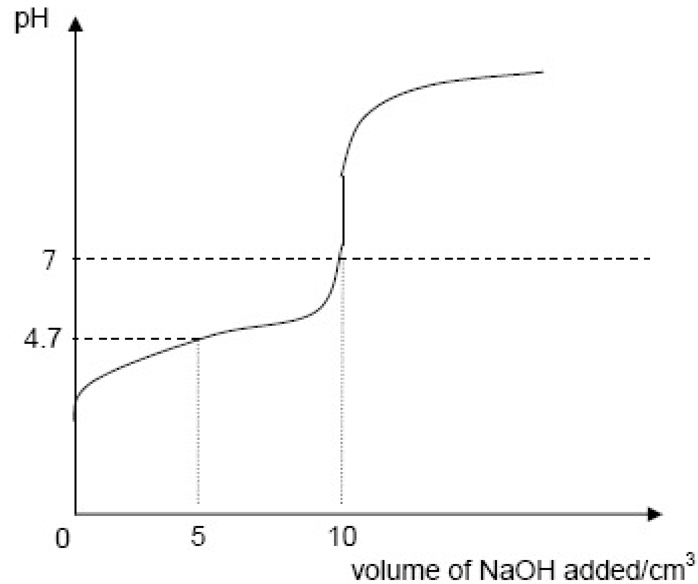
(ii) $\frac{n_{NaOH}}{n_{HA}} = 1 = \frac{c_{NaOH} \times V_{NaOH}}{c_{HA} \times V_{HA}}$

$\Rightarrow V_{NaOH} = \frac{c_{HA} \times V_{HA}}{c_{NaOH}} = \frac{0.100 \times 20.0 \times 10^{-3}}{0.200} = 0.0100 dm^3$ = **10.0 cm^3**

(iii) $A^- + H_2O \rightarrow HA + OH^-$, pH = 8.5 (any value from 8 to 9)

(iv) $pH = pK_a = -\log(2.05 \times 10^{-5})$ = **4.7**

(v)



(vi) Phenolphthalein. End point pH lies within the pH range of the indicator

34 YJC/2009/P3/Q1(e)

(e) **(i)** $K_{sp} = [Mg^{2+}(aq)]\,[OH^-(aq)]^2$: units: mol^3dm^{-9}

(ii) $[Mg^{2+}(aq)]$ in the saturated solution

$= \sqrt[3]{\frac{2.00 \times 10^{-11}}{4}} = 1.71 \times 10^{-4}\, mol\, dm^{-3}$

(iii) $[Mg^{2+}(aq)]$ extracted $= 0.0540 - 1.71 \times 10^{-4} = 0.0538$ moldm^{-3}

Maximum % of Mg extracted = $\frac{0.0538}{0.0540} \times 100 = 99.6\%$