Assessment Schedule – 2007

Describe properties of aqueous systems (90700)

Evidence Statement

Q	Evidence	Achievement	Achievement with Merit	Achievement with Excellence
ONE (a)(i)	$CH_3NH_2 + H_2O \leftrightarrow CH_3NH_3^+ + OH^-$ $(NH_4Cl \rightarrow NH_4^+ + Cl^-)$ $NH_4^+ + H_2O \leftrightarrow NH_3 + H_3O^+$	ONE equation correct.		
(ii)	$CH_3NH_2 > OH^- > (\text{or} =) CH_3NH_3^+ > H_3O^+$ $Cl^- > NH_4^+ > H_3O^+ > (\text{or} =) NH_3 > OH^-$	All species for ONE solution identified.	All species for ONE solution identified, AND the order correct for one solution.	
(b)	CH_3NH_2 is a weak base and only reacts slightly with water / it is a weak base, equilibrium lies to the left/only partial dissociation. As most aminomethane remains in the molecular state there are few ions in the solution, making it a weak electrolyte.	Weak base property recognised / recognises the solution contains a low concentration of ions/mentions partial dissociation.	Links the weak base properties/partial dissociation of CH_3NH_2 to the low concentration of ions in the solution.	
TWO (a)(i)	pH greater than 7.	pH greater than 7 with valid reason. Salt of a weak acid/Cit^{3-} accepts a proton from water/Cit^{3-} is a weak base/Cit^{3-} forms OH^- ions when it dissolves in water.	As for achieved PLUS A correct balanced equation OR Discussion of increase of OH^- concentration causing rise in pH.	All points referred to in Merit.
(a)(ii)	Since citric acid is a weak acid the anion is weakly basic and will react with water to accept H^+ from water. eg $Cit^{3-} + H_2O \leftrightarrow HCit^{2-} + OH^-$ Increase in $[OH^-]$ causes increase in pH.			
(b)(i)	$Mg(OH)_2(s) \leftrightarrow Mg^{2+}(aq) + 2\ OH^-(aq)$ $K_s = 4s^3 = 1.25 \times 10^{-11}$ $s^3 = 3.125 \times 10^{-12}$ $s = 1.46 \times 10^{-4}$	Recognises $4s^3$ relationship.	Solubility correct .	

	solubility is 1.46×10^{-4} mol L^{-1}			
THREE	$NH_3(aq) + H_2O \leftrightarrow NH_4^+(aq) + OH^-(aq)$ As NH_4Cl dissolves $[NH_4^+]$ is increased. This causes the equilibrium to move to favour formation of reactants so that $[OH^-]$ is decreased. As $[OH^-]$ is decreased, $[H_3O^+]$ is increased and pH is decreased.	Recognises $[NH_4^+]$ is increased/correct equation/Addition of $[NH_4^+]$ increases conc of H_3O^+ decreasing pH.	ONE aspect of full discussion omitted. Must include NH_3 equation,NH_4^+ addition shifts equil to left,drop in OH^- concentration. eg no link between pH and $[H_3O^+]/[OH^-]$.	Full discussion including equation.
FOUR (a)(i)	$n(NaOH) = \frac{40}{1\,000} \times 0.160 = 6.40 \times 10^{-3}$ mol $n(HPr) = 6.40 \times 10^{-3}$ mol $c(HPr) = \frac{1\,000}{50} \times 6.40 \times 10^{-3} = 0.128$ mol L^{-1}	Concentration correctly calculated.		
(ii)	$K_a = 1.35 \times 10^{-5} = \frac{[H_3O^+]\,[Pr^-]}{[HPr]}$ $[H_3O^+]^2 = 1.35 \times 10^{-5} \times 0.128 = 1.728 \times 10^{-6}$ $[H_3O^+] = 1.32 \times 10^{-3}$ mol L^{-1} pH = 2.88	pH incorrect but some processing correct, eg correctly substituted K_a expression.	pH correctly calculated.	

(b)	$Pr^- + H_2O \leftrightarrow HPr + OH^-$ $[Pr^-] = 6.40 \times 10^{-3} \times \frac{1000}{90} = 0.0711 \text{ mol L}^{-1}$ $K_a = 1.35 \times 10^{-5} = \frac{[H_3O^+][Pr^-]}{[HPr]} = \frac{[H_3O^+] \times 0.0711 \times [H_3O^+]}{10^{-14}}$ $[H_3O^+]^2 = \frac{1.35 \times 10^{-5} \times 10^{-14}}{0.0711} = 1.90 \times 10^{-18}$ $[H_3O^+] = 1.38 \times 10^{-9} \text{ mol L}^{-1}$ pH = 8.86 **OR** $K_b = [OH^-]^2 / [Pr^-] = K_w/K_a$ $[Pr^-] = 0.128 \times 50/90 = 0.711 \text{molL}^{-1}$ $[OH^-] = \sqrt{1 \times 10^{-14} \times 0.711 / 1.35 \times 10^{-5}} = 7.25 \times 10^{-6} \text{ molL}^{-1}$ pOH = 5.14 pH = 8.86	pH incorrect but some processing correct. Eg: solution recognised as basic either by writing the balanced equation OR writing a correct K_b expression OR by writing a correct formula / correctly substituted formula to use in the calculation OR correct $[Pr^-]$ calculated.	Method for calculation of pH generally correct with only one error. Eg: concentration of Pr^- incorrect.	pH correctly calculated.
(c)(i)	A buffer is a solution that undergoes minimal change of pH when small amounts of strong acid or base are added. In order to do this the buffer must contain species capable of reacting with the added acid or base.	Function of a buffer solution recognised.	Discussion shows good understanding of this buffer with at least one correct equation showing reaction with acid and base.	Full discussion of this buffer, including equations, and the limitations of the solution as a buffer.
(c)(ii)	In this system added acid would need to react with Pr^- ions $Pr^- + H_3O^+ \rightarrow HPr + H_2O$ and added base would need to react with HPr $HPr + OH^- \rightarrow Pr^- + H_2O$ Since 40 mL of sodium hydroxide is needed to exactly react with the propanoic acid, when 35 mL have been added, most of the propanoic acid has reacted and formed Pr^-. So although the solution will buffer added acid it will not be very effective if base is added. So the solution would not be a good buffer.			
(d)	Thymol blue or phenolphthalein are the best indicators because the pH range for the colour change will be approx 8.9 ± 1. This lies within the range of the equivalence point ie the vertical portion of the graph / the portion where there is a large change in pH.	Chooses both suitable indicators OR chooses only one of the two suitable indicators with a valid reason.	Thymol blue AND phenolphthalein selected and linked to the pH range for the colour change and the pK_a values.	

Judgement Statement

Achievement	Achievement with Merit	Achievement with Excellence
SIX opportunities answered at Achievement level (or higher). Minimum of 6 × A	SEVEN opportunities answered including at least FIVE at Merit level (or higher) and TWO at Achievement level (or higher). Minimum 5 × M + 2 × A	EIGHT opportunities answered including at least TWO at Excellence level plus FIVE at Merit level (or higher) and ONE at Achievement level (or higher). Minimum 2 × E + 5 × M + 1 × A

NOTE: Question Two (b) (ii) did not contribute to the overall judgement.