Assessment Schedule – 2019

Chemistry: Demonstrate understanding of equilibrium principles in aqueous systems (91392)

Evidence Statement

Q	Evidence	Achievement	Merit	Excellence
ONE (a)(i)	$Zn(OH)_2 \rightleftharpoons Zn^{2+} + 2OH^-$	Correct equilibrium equation and K_s expression.		
(ii)	$K_s = [Zn^{2+}][OH^-]^2$			
(iii)	Let solubility be 's': $[Zn^{2+}] = s$ $[OH^-] = 2s$ $K_s = 4s^3$ $4s^3 = 3.80 \times 10^{-17}$ $s = 2.12 \times 10^{-6}$ mol L^{-1} $[Zn^{2+}] = 2.12 \times 10^{-6}$ mol L^{-1} $[OH^-] = 4.24 \times 10^{-6}$ mol L^{-1}	Method correct for determining solubility.	Correct solubility of $Zn(OH)_2$ calculated, including $[Zn^{2+}]$ and $[OH^-]$.	
(iv)	$K_s = [Zn^{2+}][OH^-]^2$ $3.80 \times 10^{-17} = \left(\frac{0.210}{2}\right) \times [OH^-]^2$ $[OH^-] = 1.90 \times 10^{-8}$ mol L^{-1} Note: $\frac{0.210}{2}$ due to dilution factor from adding two solutions together.	Recognises Zn^{2+} is the common ion. OR ONE step correct. E.g. correct substitution into K_s excluding dilution / 0.105.	Calculates $[OH^-]$, but does not take into account the dilution factor.	Calculates correct $[OH^-]$. Including sig fig and units
(b)	$Zn(OH)_2 \rightleftharpoons Zn^{2+} + 2OH^-$ When the OH^- ions are in excess, the Zn^{2+} ions can form a complex ion: $Zn^{2+} + 4OH^- \rightarrow [Zn(OH)_4]^{2-}$ A decrease in $[Zn^{2+}]$ will result in the forward reaction being favoured, to restore equilibrium / minimise the change. This causes more solid $Zn(OH)_2$ to dissolve / the solubility of $Zn(OH)_2$ increases / so that $[Zn^{2+}][OH^-]$ will again equal K_s.	Recognises solubility increases due to formation of complex ion.	Explains that formation of a complex ion will decrease $[Zn^{2+}]$.	Fully explains, using equilibrium principles, how the solubility of the $Zn(OH)_2$ increases when excess NaOH is added. Must include complex ion equation.

(c)	$[Zn^{2+}] = \frac{20}{50} \times 0.0242 = 9.68 \times 10^{-3}$ mol L^{-1} $[OH^-] = \frac{30}{50} \times 1{\times}10^{-14}/10^{-13.1} = 0.0755$ mol L^{-1} IP = $[Zn^{2+}][OH^-]^2 = 9.68 \times 10^{-3} \times (0.0755)^2$ $= 5.52 \times 10^{-5}$ (5.53×10^{-5}) Since IP > K_s, a precipitate of $Zn(OH)_2$ will form.	Correct substitution into Q_s (IP) expression. OR Correct $[Zn^{2+}]$ or $[OH^-]$.	Correct process to determine Q_s and compare with K_s.	Correct calculation and comparison with K_s to determine whether $Zn(OH)_2$ will form a precipitate.

NØ	N1	N2	A3	A4	M5	M6	E7	E8
No response; no relevant evidence.	1a	2a	3a	4a	2m	3m	1e	2e

Q	Evidence	Achievement	Merit	Excellence
TWO (a)(i)	Na^+, $HCOO^-$, HCOOH, OH^-, H_3O^+	THREE species identified.		
(ii)	After 12.5 mL NaOH has been added, it is halfway to the equivalence point. This means that [HCOOH] = [$HCOO^-$]. Therefore pH equals pK_a.	[HCOOH] = [$HCOO^-$] OR pH = pK_a	[HCOOH] = [$HCOO^-$] Therefore pH = pK_a.	
(b)(i)	Tick cresol red. Indicators change colour at a pH ±1 of the pK_a / near the pK_a. Therefore, cresol red should be used as it will change near the equivalence point/steepest part of the curve, whereas thymol blue and bromocresol green will change before the equivalence point/steepest part of the curve.	Identifies cresol red. OR Recognises an indicator is chosen to change colour over the vertical section of the curve / at the equivalence point.	Explanation with reference to the pK_a of why cresol red is the appropriate indicator whereas the other two indicators are not.	Full explanation for indicator choice . AND pH at equivalence point.
(ii)	$\left[H_3O^+\right]=\sqrt{\frac{K_a \times K_w}{\left[CH_3CH_2NH_2\right]}}$ $\left[H_3O^+\right]=\sqrt{\frac{1.82\times10^{-4}\times1\times10^{-14}}{0.0778}}$ $\left[H_3O^+\right]=4.84\times10^{-9}$ mol L^{-1} $pH=-\log\left[H_3O^+\right]=8.32$ OR For this solution, $HCOO^- + H_2O \rightleftharpoons HCOOH + OH^-$	Correct process for determining pH at equivalence point. OR One correct step.	pH calculated at equivalence point with incorrect dilution.	
(c)	After 28 mL NaOH added: n(unreacted NaOH) = cV = 0.140 × 0.003 = 4.2 × 10^{-4} mol $c(\text{NaOH})=\frac{n}{V}=\frac{4.2\times10^{-4}}{0.048}=8.75\times10^{-3}$ mol L^{-1} $\left[H_3O^+\right]=\frac{1\times10^{-14}}{8.75\times10^{-3}}=1.14\times10^{-12}$ mol L^{-1} $pH=-\log1.14\times10^{-12}=11.9$	Correct n(NaOH). OR One correct step.	Correct process but one error in calculation.	Correct pH.

NØ	N1	N2	A3	A4	M5	M6	E7	E8
No response; no relevant evidence.	1a	2a	3a	4a	2m	3m	1e	2e

Q	Evidence	Achievement	Merit	Excellence
THREE (a)(i)	CH_3COOH has the smaller pK_a / larger K_a, so it will be a stronger acid than NH_4^+. This means CH_3COOH will dissociate to a greater extent to produce a higher $[H_3O^+]$ than NH_4^+, so CH_3COOH will have a lower pH.	Recognises pK_a is a measure of the extent of dissociation of a weak acid.	Links magnitude of pK_a (or K_a) to degree of dissociation, $[H_3O^+]$ and pH. OR	Full explanation of the pH and electrical conductivity of BOTH solutions, i.e. links degree of dissociation to BOTH $[H_3O^+]$ and pH AND [ions] and electrical conductivity, including supporting equations.
(ii)	The electrical conductivity of a solution depends upon [ions]. NH_4Cl is an acidic salt, so it completely dissociates / dissolves to produce a relatively high $[NH_4^+]$ and $[Cl^-]$. This makes NH_4Cl a good electrical conductor. $NH_4Cl \rightarrow NH_4^+ + Cl^-$ However, CH_3COOH is a weak acid and only partially dissociates to produce a relatively low $[CH_3COO^-]$ and $[H_3O^+]$. This makes CH_3COOH a poorer electrical conductor than NH_4Cl. $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$	Recognises electrical conductivity in a solution requires ions.	Links the [ions] to the electrical conductivity of BOTH solutions.	
(iii)	$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$ $K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$ $10^{-4.76} = \frac{(1.78 \times 10^{-3})^2}{[CH_3COOH]}$ $[CH_3COOH] = 0.182 \text{ mol L}^{-1}$	Correct process for determining $[CH_3COOH]$. (*correct substitution into formula)* OR One correct step	Correct $[CH_3COOH]$.	

(b)(i)	$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$ $K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$ $10^{-4.76} = 2 \times \frac{[H_3O^+]}{5}$ $[H_3O^+] = 4.34 \times 10^{-5}$ mol L^{-1} $pH = -\log 4.34 \times 10^{-5} = 4.36$	Correct process for determining pH. OR One correct step.	Correct pH.	
(ii)	$[CH_3COOH] > [CH_3COO^-]$ / $pH < pK_a$ Therefore the buffer solution is more effective at neutralising strong base: $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$	Equation. OR $[CH_3COOH] > [CH_3COO^-]$	Links ratio of $[CH_3COOH]$:$[CH_3COO^-]$ to buffer effectiveness, including equation.	Full explanation of buffer behaviour for (ii) and (iii).
(iii)	When water is added, the ratio of CH_3COOH to CH_3COO^- is unchanged, so the pH of the buffer solution is unaffected.	Recognises pH remains unchanged.	Explains effect of dilution on pH in terms of ratio between CH_3COOH and CH_3COO^-.	

NØ	N1	N2	A3	A4	M5	M6	E7	E8
No response; no relevant evidence.	1a	2a	3a	4a	2m	3m	1e	2e

Cut Scores

Not Achieved	Achievement	Achievement with Merit	Achievement with Excellence
0 – 6	7 – 13	14 – 19	20 – 24