Assessment Schedule – 2014

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

Evidence Statement

Q	Evidence	Achievement	Achievement with Merit	Achievement with Excellence
ONE (a)(i) (ii)	$HOCl + H_2O \rightleftharpoons OCl^- + H_3O^+$ $HOCl > H_3O^+ > OCl^- > OH^-$ or $HOCl > H_3O^+ = OCl^- > OH^-$ HOCl partially dissociates, and so the equilibrium lies to the LHS/favours the reactants; therefore HOCl is present in the greatest amounts. H_3O^+ and OCl^- are produced in equal amounts / there is a small contribution to H_3O^+ from water therefore $H_3O^+ > OCl^-$ Because there is a relatively high $[H_3O^+]$, the $[OH^-]$ is very low (or links to K_w).	• Equation correct. OR FOUR species correctly identified. • Recognises HOCl partially dissociates. OR One correct justification.	• ALL species and order correct AND partial explanation to support the order of the species.	• ALL species and order correct AND complete justification.
(b)	Hydrofluoric acid is a stronger acid/more acidic/dissociates more because it has a smaller pK_a (larger K_a) than hypochlorous acid. So HF will therefore have a higher $[H_3O^+]$. As $[H_3O^+]$ increases, the pH decreases, so HF will have a lower pH than HOCl. (pH HF = 2.09, HOCl = 4.27)	• Any two correct relationships.	• Complete comparison.	
(c)	$K_a = \frac{[F^-][H_3O^+]}{[HF]}$ $10^{-3.17} = \frac{[F^-] \times 10^{-4.02}}{0.0500}$ $[F^-] = 0.354 \text{ mol L}^{-1}$ $n(NaF) = 0.354 \text{ mol L}^{-1} \times 0.150 \text{ L} = 0.0531 \text{ mol}$ $m(NaF) = 0.0531 \text{ mol} \times 42.0 \text{ g mol}^{-1} = 2.23 \text{ g}$	• Writes correct K_a or pH expression. OR Calculates K_a or $[H_3O^+]$. • Correct 'n' and 'm' step with incorrect $[F^-]$.	• Correct method but error in calculation / units missing / unit incorrect.	• Correct answer with units.

NØ	N1	N2	A3	A4	M5	M6	E7	E8
No response or no relevant evidence.	1a	2a	3a	4a	2m	3m	2e with minor error / omission.	2e

Q	Evidence	Achievement	Achievement with Merit	Achievement with Excellence
TWO (a)(i) (ii)	$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$ $K_s = [Pb^{2+}][Cl^-]^2$	• Both (i) and (ii) correct.		
(iii)	$[Pb^{2+}] = x \quad [Cl^-] = 2x$ $K_s = 4x^3$ $x = \sqrt[3]{\frac{K_s}{4}}$ $= \sqrt[3]{\frac{1.70 \times 10^{-5}}{4}}$ $= 1.62 \times 10^{-2}\ \text{mol L}^{-1}$ $[Pb^{2+}] = 1.62 \times 10^{-2}\ \text{mol L}^{-1}$ $[Cl^-] = 3.24 \times 10^{-2}\ \text{mol L}^{-1}$	• Method correct, for calculation of solubility.	• Correct answer for solubility and both $[Pb^{2+}]$ and $[Cl^-]$.	
(b)	$n(Pb(NO_3)_2) = \frac{2.00\ \text{g}}{331\ \text{g mol}^{-1}}$ $= 6.04 \times 10^{-3}$ mol $\therefore [Pb^{2+}] = 6.04 \times 10^{-3}$ mol / 0.500L $= 1.21 \times 10^{-2}\ \text{mol L}^{-1}$ $Q = (1.21 \times 10^{-2})$ x $(0.440)^2$ $= 2.34 \times 10^{-3}$ As $Q > K_s$, **a precipitate will form.**	• One calculation step correct. • Compares incorrect Q and K_s to make a valid conclusion.	• One calculation error AND Compares Q and K_s to make a valid conclusion.	• Answer correct with supporting calculation and correct conclusion.

(c)	$Zn(OH)_2(s) \rightleftharpoons Zn^{2+}(aq) + 2OH^-(aq)$ When pH is less than 4 / low, $[OH^-]$ is decreased due to the reaction with H_3O^+ to form water, $H_3O^+ + OH^- \rightarrow H_2O$ so equilibrium shifts to the right to produce more $[OH^-]$, therefore more $Zn(OH)_2$ will dissolve. When pH is greater than 10 / high, then more OH^- is available and the complex ion (zincate ion) will form. $Zn(OH)_2(s) + 2OH^- \rightarrow [Zn(OH)_4]^{2-}$ OR $Zn^{2+} + 4OH^- \rightarrow [Zn(OH)_4]^{2-}$ This decrease in $[Zn^{2+}]$ causes the position of equilibrium to shift further to the right, therefore more $Zn(OH)_2$ dissolves.	• Writes the equilibrium equation. • Recognises solubility increases at pH of less than 4 (acidic conditions) **due to removal of OH^-.** OR Recognises the solubility increases at a pH greater than 10 **due to formation of a complex ion.**	• Partial explanation for BOTH changes in pH, not fully related to the effect on the equilibrium. OR One change in pH fully explained.	• Complete explanation for BOTH changes in pH.

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No response or no relevant evidence.	1a	2a	3a	4a	2m	3m	2e with minor error / omission.	2e

Q	Evidence	Achievement	Achievement with Merit	Achievement with Excellence
THREE (a)	At point **A**, $[CH_3NH_2] \approx [CH_3NH_3^+]$. So the solution has buffering properties in the proximity of point **A**. When HBr is added, the H_3O^+ is consumed: $H_3O^+ + CH_3NH_2 \rightarrow CH_3NH_3^+ + H_2O$ Since the H_3O^+ is removed from the solution (neutralised), the pH does not change significantly.	• Recognises near point **A** solution is a buffer / $[CH_3NH_2] \approx [CH_3NH_3^+]$. • Identifies H_3O^+ or HBr is neutralised / removed by CH_3NH_2. • Calculates $[OH^-]$ / $[H_3O^+]$ / K_b	• Correct equation linked to neutralisation / absorption of H_3O^+.	
(b)	$[H_3O^+] = 10^{-11.8} = 1.58 \text{ x } 10^{-12}$ $K_a = \frac{[CH_3NH_2][H_3O^+]}{[CH_3NH_3^+]}$ $= \frac{[CH_3NH_2][H_3O^+]}{[OH^-]}$ $2.29 \times 10^{-11} = \frac{[CH_3NH_2] \times (10^{-11.8})^2}{1 \times 10^{-14}}$ $[CH_3NH_2] = \frac{(2.29 \times 10^{-11}) \times (1 \times 10^{-14})}{(10^{-11.8})^2}$ $= 0.0912 \text{ mol L}^{-1}$ **OR** $[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{10^{-14}}{10^{-11.8}}$ $= 6.31 \times 10^{-3} \text{ mol L}^{-1}$ $K_b = \frac{[OH^-]^2}{[CH_3NH_2]}$ $4.37 \times 10^{-4} = \frac{(6.31 \times 10^{-3})^2}{[CH_3NH_2]}$ $[CH_3NH_2] = \frac{(6.31 \times 10^{-3})^2}{4.37 \times 10^{-4}}$ $[CH_3NH_2] = 0.0912 \text{ mol L}^{-1}$	• Uses suitable process with more than one error. OR Rearranges K_b / K_a expression so $[CH_3NH_2]$ is the subject.	• Correct method but an error in the calculation.	• Correct answer with a clear method.

(c)(i)	$CH_3NH_3^+$, Br^-, CH_3NH_2, H_3O^+	TWO OF: • ALL species correct. • Recognises ions are required for electrical conductivity in a solution. • One correct equation.	• Full explanation of the electrical conductivity and species present of either the initial CH_3NH_2 solution or the solution at point B. OR for an answer discussing each solution separately: TWO OF: Species and comparative concentrations within each solution for both solutions / two of the three equations / conductivity of each solution with reasons.	• Compares and contrasts the electrical conductivity of BOTH the initial CH_3NH_2 solution and the solution at point **B**, including a consideration of the differing concentrations of each solution.
(ii)	**At the start**, before addition of HBr there is a solution of weak base (CH_3NH_2) which only partially reacts with water to produce a relatively low concentration of ions. As a result, the initial CH_3NH_2 solution will be a poor electrical conductor. $\mathbf{CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-}$ Therefore species present are $CH_3NH_2 > OH^- \geq CH_3NH_3^+ > H_3O^+$ **At point B,** there is a solution of the salt CH_3NH_3Br present which is dissociated completely into ions. Therefore there is a relatively high concentration of ions ($CH_3NH_3^+$ and Br^-) present in the solution, so it will be a good electrical conductor / electrolyte. $\mathbf{CH_3NH_3Br \rightarrow CH_3NH_3^+ + Br^-}$ $CH_3NH_3^+$ reacts with water according to the equation $\mathbf{CH_3NH_3^+ + H_2O \rightleftharpoons CH_3NH_2 + H_3O^+}$ Species present are $Br^- > CH_3NH_3^+ > H_3O^+ \geq CH_3NH_2 > (OH^-)$			

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Cut Scores

	Not Achieved	Achievement	Achievement with Merit	Achievement with Excellence
Score range	0 – 7	8 – 13	14 – 18	19 – 24