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)	$K_a = \dfrac{[F^-][H_3O^+]}{[HF]}$ $10^{-3.17} = \dfrac{[F^-] \times 10^{-4.02}}{0.0500}$ $[F^-] = 0.354 \text{ mol L}^{-1}$ $n(NaF) = 0.354 \text{ mol L}^{-1} \times 0.150 \text{ L} = 0.0531 \text{ mol}$ $m(NaF) = 0.0531 \text{ mol} \times 42.0 \text{ g mol}^{-1} = 2.23 \text{ g}$	• 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)	$[Pb^{2+}] = x \quad [Cl^-] = 2x$ $K_s = 4x^3$ $x = \sqrt[3]{\frac{K_s}{4}}$ $= \sqrt[3]{\frac{1.70 \times 10^{-5}}{4}}$ $= 1.62 \times 10^{-2}$ mol L^{-1} $[Pb^{2+}] = 1.62 \times 10^{-2}$ mol L^{-1} $[Cl^-] = 3.24 \times 10^{-2}$ 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 $= 1.21 \times 10^{-2}$ mol L^{-1} $Q = (1.21 \times 10^{-2}) \times (0.440)^2$ $= 2.34 \times 10^{-3}$ 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.

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
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 • Uses suitable process with more than one error. OR Rearranges K_b / K_a expression so $[CH_3NH_2]$ is the subject.	• Correct equation linked to neutralisation / absorption of H_3O^+. • Correct method but an error in the calculation.	• Correct answer with a clear method.
(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** $[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{10^{-14}}{10^{-11.8}}$ $= 6.31 \times 10^{-3} \text{ mol L}^{-1}$ $K_b = \frac{[OH^-]^2}{[CH_3NH_2]}$ $4.37 \times 10^{-4} = \frac{(6.31 \times 10^{-3})^2}{[CH_3NH_2]}$ $[CH_3NH_2] = \frac{(6.31 \times 10^{-3})^2}{4.37 \times 10^{-4}}$ $[CH_3NH_2] = 0.0912 \text{ mol L}^{-1}$			

(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. **$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$** 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. **$CH_3NH_3Br \rightarrow CH_3NH_3^+ + Br^-$** $CH_3NH_3^+$ reacts with water according to the equation **$CH_3NH_3^+ + H_2O \rightleftharpoons CH_3NH_2 + H_3O^+$** Species present are $Br^- > CH_3NH_3^+ > H_3O^+ \geq CH_3NH_2 > (OH^-)$			

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

Cut Scores

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