Assessment Schedule – 2015

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

Evidence Statement

Q	Evidence	Achievement	Achievement with Merit	Achievement with Excellence
ONE (a)(i)	$CH_3NH_3Cl(s) \xrightarrow{H_2O} CH_3NH_3^+(aq) + Cl^-(aq)$	• One correct equation. OR Four species identified.		
(ii)	$CH_3NH_3^+ + H_2O \rightleftharpoons CH_3NH_2 + H_3O^+$			
(iii)	$Cl^- > CH_3NH_3^+ > H_3O^+ = CH_3NH_2 > OH^-$ OR $Cl^- > CH_3NH_3^+ > H_3O^+ > CH_3NH_2 > OH^-$		• Correct order, all species.	
(iv)	$K_a = \dfrac{[CH_3NH_2][H_3O^+]}{[CH_3NH_3^+]}$ $2.29 \times 10^{-11} = \dfrac{[H_3O^+]^2}{0.0152}$ $[H_3O^+] = 5.90 \times 10^{-7}$ mol L^{-1} pH = –log $5.90 \times 10^{-7} = 6.23$ $\boxed{\begin{array}{l}\left[H_3O^+\right] = \sqrt{K_a} \times HA \\ \quad = 5.90 \times 10^{-7} \\ pH = 6.23\end{array}}$	• Correct process.	• Correct pH.	

(b)	**pH:** The pH of a solution is calculated from its [H_3O^+]. NaOH is an ionic solid that is a strong base and dissociates completely to produce a high OH^- concentration (low [H_3O^+]). Since [OH^-] is high / [H_3O^+] is low, the pH is high. $NaOH \rightarrow Na^+ + OH^-$ CH_3NH_2 is a weak base that partially reacts / dissociates / ionises with H_2O producing a lower concentration of OH^-, Therefore it has a lower pH than NaOH: $CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$ The CH_3COONa is an ionic solid that dissociates completely in H_2O. The CH_3COO^- ion is a weak base that partially reacts / dissociates / ionises with H_2O producing a lower concentration of OH^-. $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$ The pH is closer to 7, showing it is the weakest base. Therefore it has a lowest pH **Electrical conductivity:** Electrical conductivity is determined by the concentration of ions. NaOH completely dissolves to produce a high concentration of Na^+ and OH^- ions in solution. $NaOH \rightarrow Na^+ + OH^-$ Therefore it is a good conductor. Since CH_3NH_2 is a weak base, it only partially reacts with water to produce a low concentration of ions in solution so it is a poor electrical conductor. $CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$ CH_3COONa is also an ionic solid. It dissolves completely to produce a high concentration of Na^+ and CH_3COO^- ions: $CH_3COONa \rightarrow Na^+ + CH_3COO^-$ Therefore it is a good conductor.	• Recognises that pH depends upon [H_3O^+] / [OH^-] ratio. OR States that NaOH is a strong base whereas CH_3COONa and CH_3NH_2 form weakly basic solutions. • Recognises that electrical conductivity depends upon concentration / amount of ions in solution. OR Relates conductivity to the degree of dissociation.	• Links the pH of each solution to its strength and degree of dissociation. • Links the ion concentration / amount of each solution to the degree of dissociation and its electrical conductivity.	• Compares and contrasts the pH of each solution. • Compares and contrasts the electrical conductivity of each solution.

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

Q	Evidence	Achievement	Achievement with Merit	Achievement with Excellence
TWO (a)(i)	$CaCO_3(s) \rightleftharpoons Ca^{2+}(aq) + CO_3^{2-}(aq)$	• (i) and (ii) correct.		
(ii)	$K_s = [Ca^{2+}][CO_3^{2-}]$			
(iii)	$K_s(CaCO_3) = (5.74 \times 10^{-5})^2 = 3.29 \times 10^{-9}$	• Method correct.	• Correct answer for $K_{s.}$	
(b)	The H_3O^+ from the acidic solution reacts with the CO_3^{2-}. This reduces $[CO_3^{2-}]$, causing the equilibrium to shift towards the products / RHS to replace some of the lost CO_3^{2-}. Therefore more solid $CaCO_3$ will dissolve. $2H_3O^+ + CO_3^{2-} \rightarrow 3H_2O + CO_2$ (or other correct alternative).	• Recognises H_3O^+ will remove / neutralise the CO_3^{2-} from the equilibrium.	• Recognises H_3O^+ will remove / reacts with CO_3^{2-} with a relevant balanced equation AND uses equilibrium principles to link to an increased solubility of $CaCO_3$.	
(c)	$Pb(OH)_2 \rightleftharpoons Pb^{2+} + 2OH^-$ $Q = [Pb^{2+}][OH^-]^2$ $[Pb^{2+}] = 0.5 \times 0.00421 = 2.105 \times 10^{-3}$ $[OH^-] = 0.5 \times 0.0398 = 1.99 \times 10^{-2}$ $Q = (2.105 \times 10^{-3}) \times (1.99 \times 10^{-2})^2$ $Q = 8.34 \times 10^{-7}$ Since $Q > K_s$, a precipitate of $Pb(OH)_2$ will form. pH =12.6 pOH = 1.4 $[OH^-] = 0.0398$	• Correct equation / expression. • Compares incorrect Q value to K_s, to prove $Pb(OH)_2$ forms a precipitate. OR Correct working.	• Method uses correct Q expression but has one calculation error AND Compares Q and K_s to prove a $Pb(OH)_2$ precipitate forms.	• Correct calculation with comparison of Q and K_s to prove $Pb(OH)_2$ forms a precipitate.

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

Q	Evidence	Achievement	Achievement with Merit	Achievement with Excellence
THREE (a)(i)	Na^+, F^-, H_2O, HF, OH^-, H_3O^+.	• Three species listed.		
(ii)	A weak base, F^-, is present at the equivalence point: $F^- + H_2O \rightleftharpoons HF + OH^-$ This increase in $[OH^-]$ causes the pH to be greater than 7.	• Recognises F^- is the weak base responsible for pH > 7. OR Correct equation.	• Links pH at equivalence to increased $[OH^-]$, including an equation for the dissociation of F^-.	
(iii)	$K_a = \frac{[F^-][H_3O^+]}{[HF]}$ $10^{-3.17} = \frac{1 \times [H_3O^+]}{2}$ $[H_3O^+] = 2 \times 10^{-3.17} = 1.35 \times 10^{-3}$ mol L^{-1} pH = –log (1.35×10^{-3}) = 2.87. $\boxed{\begin{aligned} pH &= pK_a + \log [F^-] / [HF] \\ &= 3.17 + \log 0.5 \\ &= 2.87 \end{aligned}}$ Since there are significant concentrations of the weak acid and its conjugate base the solution can resist added acid or base. However, since the pH of the buffer solution is less than the pK_a, / $[HF] > [F^-]$, it is more effective against added base than acid.	• Correct process for determining the pH. • Recognises solution is more effective against added base. OR Describes function of a buffer by resisting added acid and base.	• Correct pH OR Evaluates the function of the buffer.	• Correct pH and full evaluation.
(iv)	$n(\text{NaOH}) = cv = 0.258 \times \frac{24-20}{1000} = 1.032 \times 10^{-3}$ mol $c(\text{NaOH}) = \frac{n}{v} = \frac{1.032 \times 10^{-3}}{44/1000} = 0.0235$ mol L^{-1} $[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1 \times 10^{-14}}{0.0235} = 4.26 \times 10^{-13}$ mol L^{-1} $pH = -\log 4.26 \times 10^{-13} = 12.4$ $\boxed{\begin{aligned} [OH^-] &= 0.258 \times 4 / 44 \\ &= 0.02345 \text{ mol } L^{-1} \\ pOH &= -\log (0.02345) \\ &= 1.63 \\ pH &= 14 - pOH \\ &= 12.4 \end{aligned}}$	• Correct n(NaOH). OR One correct step.	• Correct process but one error in calculation.	• Correct answer.

(b)	Since CH_3COOH has a higher pK_a, it is a weaker acid than HF. Therefore its conjugate base, CH_3COO^-, will be a stronger base than F^-. This means $[OH^-]$ will be higher at the equivalence point for the CH_3COOH vs NaOH titration, so the equivalence point pH will be higher.	• States CH_3COOH is a weaker acid than HF. OR States equivalence point is higher with some evidence.	• Links higher equivalence point for CH_3COOH to the strength of the weak acid, its pK_a, and the relative strength of its conjugate base.	• Contrasts the equivalence point for both titrations.

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

Cut Scores

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