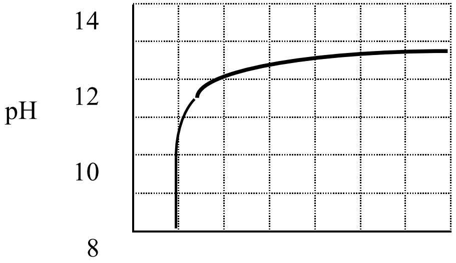
Assessment Schedule – 2008

Chemistry: Describe properties of aqueous systems (90700)

Evidence Statement

Q	Evidence	Achievement	Achievement with Merit	Achievement with Excellence
ONE (a)(i)	X drawn **on curve** at pH 9 (accept 8.5–9.5).	pH or graph correct. Graph line must go above 12 and not touch 13. Line must also not trend upwards at the end.	Both correct.	
(ii)	End of graph is drawn towards an asymptote line at pH 13. 			
(b)	$HN_3(aq) + H_2O \rightleftharpoons H_3O^+(aq) + N_3^-(aq)$ $K_a(CH_3COOH) = \frac{[H_3O^+][N_3^-]}{[HN_3]}$ pH = 2.6 so $[H_3O^+(aq)] = \mathbf{2.51 \times 10^{-3}\ mol\ L^{-1}} = [N_3^-(aq)]$ $pK_a(HN_3) = 4.72$, $K_a(HN_3) = \mathbf{1.91 \times 10^{-5}}$ $K_a(HN_3) = 1.91 \times 10^{-5} = \frac{(2.51 \times 10^{-3})^2}{x}$ $\boldsymbol{x = 0.331}$ **ie** $\mathbf{[HN_3\,(aq)] = 0.331\ mol\ L^{-1}}$ (Allow 0.330 and 0.33)	$[H_3O^+]$ determined from pH OR K_a determined from pK_a. (As long as solution is identified.)	Concentration of acid solution correctly determined. (Must have correct unit)	

TWO (a)(i)	$HOBr(aq) + H_2O \rightleftharpoons H_3O^+(aq) + OBr^-(aq)$	Equation and K_a expression correct. Equation must have equilibrium arrow, states not required. K_a expression must have square brackets.		
(ii)	$K_a(HOBr) = \frac{[H_3O^+][OBr^-]}{[HOBr]}$			
(b)	$K_a(HOBr) = \frac{[H_3O^+][OBr^-]}{[HOBr]} = 10^{-8.62} = \mathbf{2.40 \times 10^{-9}}$ $[H_3O^+(aq)] = [OBr^-(aq)] = x \text{ mol L}^{-1}$ $\frac{x^2}{0.0525} = 2.40 \times 10^{-9} \quad x = [H_3O^+(aq)] = 1.12 \times 10^{-5}$ **pH = 4.95** (allow 4.9 or 4.950)	Method correct but minor arithmetical error. OR K_a determined from pK_a.	pH calculated correctly.	
THREE (a) (i)	$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2\,Cl^-(aq)$	Equation and K_s correct. States do not need to be shown. Square brackets required. Equilibrium arrow not required.		
(ii)	$K_s(PbCl_2) = [Pb^{2+}][Cl^-]^2$			
(b)	Let s be the solubility $K_s = s(2s)^2$ $1.60 \times 10^{-5} = 4s^3$ $s = 0.0159 \text{ mol L}^{-1}$ solubility of $PbCl_2(s)$ **= 0.0159 mol L^{-1}** (allow 0.0160 or 0.016)	**Method correct** but minor arithmetical error e.g cube root done before divided by 4 Or rounding error, eg 0.0158	Solubility calculated correctly. Units not required.	
(c)	$n(Pb(NO_3)_2) = \frac{1.00 \text{ g}}{331 \text{ g mol}^{-1}} = 3.02 \times 10^{-3} \text{ mol}$ If the lead nitrate is dissolved in the 500 mL, then **[Pb^{2+}] = 6.04 × 10^{-3} mol L^{-1}** $Q_s = [Pb^{2+}][Cl^-]^2 = 6.04 \times 10^{-3} \times (0.440)^2 = \mathbf{1.17 \times 10^{-3}}$ As Q_s (1.17×10^{-3}) exceeds K_s (1.60×10^{-5}), **a precipitate will form.**	ONE step correct. Eg: Correct [Pb^{2+}]. Correct solubility calculation from incorrect inputs. Correct conclusion from incorrect inputs.	**Method uses correct IP expression** but has one error. Eg [Pb^{2+}] incorrect or fails to square 0.440. Correct conclusion from calculated answer. Incorrect or no conclusion from correct calculation.	Answer correct with supporting calculation and reasoning.

FOUR	$Al(OH)_3(s) \rightleftharpoons Al^{3+}(aq) + 3\ OH^-(aq)$ **Acidic conditions** When the pH is less than 4, $[H_3O^+] > [OH^-]$ and the **$[H_3O^+]$ ions react with OH^-** ions from $Al(OH)_3$. ie $H_3O^+ + OH^- \rightarrow 2H_2O$ **(OR** $Al(OH)_3(s) + 3\ H_3O^+ \rightarrow Al^{3+}(aq) + 6\ H_2O$) The decrease in $[OH^-]$ from the solubility equilibrium causes the position of equilibrium to shift further to the right so that more $Al(OH)_3$ is dissolved. **Basic conditions** When the pH is greater than 10, $[H_3O^+] < [OH^-]$ and the OH^- ions react with Al^{3+} ions from $[Al(OH)_4]^-$ ie $Al^{3+} + 4\ OH^- \rightarrow [Al(OH)_4]^-$ **(OR** $Al(OH)_3(s) + OH^- \rightarrow [Al(OH)_4]^-$) The decrease in $[Al^{3+}]$ from the solubility equilibrium causes the position of equilibrium to shift further to the right so that more $Al(OH)_3$ is dissolved.	Any TWO of Correct K_s equation low pH $[H_3O^+] > [OH^-]$ high pH $[H_3O^+] < [OH^-]$ acidic conditions H_3O^+ removes OH^- basic conditions OH^- removes Al^{3+} / forms $[Al(OH)_4]^-$.	Solubility in either **acidic or basic conditions** explained **with** all relevant equations (Must say acidic conditions H_3O^+ removes OH^-) OR Solubility in both **acidic and basic conditions** explained **without** all equations.	Full discussion including equations, with equations related to solubility.

FIVE (a) (i) (ii)	$K_a(NH_4^+) = \frac{[H_3O^+]\,[NH_3]}{[NH_4^+]}$ so $[H_3O^+] = K_a(NH_4^+) \times \frac{[NH_4^+]}{[NH_3]}$ **Solution A** **By calculation** $[NH_4^+] = [NH_3] = 50/100 \times 1.00 =$ **0.500 mol L^{-1}** therefore $[NH_4^+] / [NH_3] = 1$ So $[NH_3] = Ka(NH_4^+)$ ie pH = pKa = 9.24 **Or by reasoning** Since equal volumes of solutions with equal concentrations of each species are mixed. $[NH_4^+] = [NH_3]$ So, $[H_3O^+] = Ka(NH_4^+)$ ie pH = pKa = 9.24 **Solution B** Because the added HCl reacts with ammonia to form ammonium ions $NH_3(aq) + H_3O^+(aq) \rightarrow NH_4^+(aq) + H_2O$ **By calculation** nHCl added $= 0.010 \times 0.025 = 2.5 \times 10^{-4} = nNH_4^+$ nNH_3 remaining $= 0.010 \times 0.050 = 5.0 \times 10^{-4} - 2.5 \times 10^{-4} = 2.5 \times 10^{-4}$ so $[NH_4^+] = [NH_3] = 2.5 \times 10^{-4} /0.075 = 3.33 \times 10^{-3}$ **(NOT $[NH_3] = 6.67 \times 10^{-3}$)** So, $[H_3O^+] = K_a (NH_4^+)$ ie $pH = pK_a = 9.24$ **Or by reasoning** Since the volume of HCl added is half the volume of NH_3 and the concentrations are the same, $n(H_3O^+)$ added equals $n(NH_4^+)$ formed and also equals $n(NH_3)$ remaining, therefore $[NH_4^+] = [NH_3]$ So, $[H_3O^+] = K_a (NH_4^+)$ ie $pH = pK_a = 9.24$	Correct process for calculation of pH of ONE solution clearly shown but answer is incorrect OR Ka expression correct **and** answer attempted with some correct information. Eg the formula for finding $[H_3O^+]$ OR pH OR a concentration correct OR calculating Ka = (5.75×10^{-10})	Correct pH of ONE solution either by calculation or explanation.	Correct pH of BOTH solutions.
(b)	(i) Both solutions A and B can act as buffers because both contain reasonable amounts of a weak acid and its conjugate base/ NH_4^+ and NH_3. The ammonium ion can absorb added base and the ammonia can absorb added acid (or equations) thereby minimising changes in pH. (ii) Sodium hydroxide is a strong base and will accept H^+ from the NH_4^+ ions in solution. $OH^- + NH_4^+ \rightarrow NH_3 + H_2O.$ and so the pH will only increase very slightly as the effect of added OH^- ions is minimised. The lower amount / concentration / number of NH_4^+ ions in solution B limit the buffering properties as these would be used up more rapidly, so the pH will increase / change faster than with Solution A. Accept evidence for part (i) and (ii) from either section.	Recognises that a buffer solution exists in A and B (can be implied) OR recognises A or B is a buffer with some more detail. eg contains a weak acid and its conjugate base resists changes in pH gives the correct equation.	Recognises A and B are buffers. Identifies the presence of a weak acid and its conjugate base by name or formula in both A and B. Explains the function of a buffer.	Effect of adding the strong base discussed for each solution. Answer includes comparison of buffering effect of solutions A and B. Correct equation included for buffer solutions.

Judgement Statement

Achievement	Achievement with Merit	Achievement with Excellence
Total of SIX opportunities answered at Achievement level or higher. 6 × A	Total of at least SEVEN opportunities answered with FOUR at Merit level or higher. 4 × M + 3 × A	Total of at least SEVEN opportunities answered with TWO at Excellence level and THREE at Merit level or higher. 2 × E + 3 × M + 2 × A