ANSWERS: NCEA exam questions on titration curves

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| **2017** | **Evidence** | **Achievement** | **Merit** | **Excellence** |
| (a)  (b) (i)  (ii)  (c) | Phenolphthalein.  CH3COOH + H2O ⇌ CH3COO– + H3O+    *n*(NaOH) added = 0.112 mol L–1 × 0.01 L = 1.12 × 10–3 mol  Initial *n*(CH3COOH) = 0.166 mol L–1 × 0.02 L = 3.32 × 10–3 mol  *n*(CH3COOH) remaining in 30 mL  = 3.32 × 10–3 mol – 1.12 × 10–3 mol  = 2.2 × 10–3 mol  *n*(CH3COO–) in 30 mL = 1.12 × 10–3 mol    CH3COO–, Na+, CH3COOH, OH–, H3O+  Both titrations produce a basic salt at the equivalence point because [OH–] > [H3O+] / pH > 7 / equations.  At the equivalence point the ethanoic acid is converted into the ethanoate ion and the methanoic acid is converted into the methanoate ion.  The methanoic acid titration has a lower pH than the ethanoic acid titration at the equivalence point because:  methanoate ions are a weaker base / produce fewer OH– ions than ethanoate ions  OR  HCOOH has a larger *K*a than CH3COOH, as it is a stronger acid. As a result, its conjugate base, HCOO–, is weaker than CH3COO, so the pH will be lower at the equivalence point. | • Correct indicator.  • Correct process with minor error.  • Correct *n*(NaOH)  OR  Initial *n*(CH3COOH).  THREE species.  Both are basic  HCOOH is a stronger acid than  CH3COOH  OR  HCOO– is weaker than CH3COO– | • Correct calculation.  • Two steps correct  Links equivalence point pH for either, to the relative strength of  the conjugate base  OR  relative concentrations of hydroxide ions. | • Correct pH, including three significant figures.  Compares and contrasts the equivalence point for both titrations. |

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| **2016** | **Evidence** | **Achievement** | **Achievement with Merit** | **Achievement with Excellence** |
| (a) | *(Ammonium chloride)* is acidic OR NH4+ + H2O  H3O+ + NH3  So therefore  [H3O+] > [OH-] | * H3O+ is produced.   OR  It is acidic. | * BOTH concepts correct. |  |
| (b) | [NH4+] = 0.320 × 20 / 30 = 0.213 molL-1  *(Ka = 10–9.24 = 5.75 × 10–10)*  [H3O+] = √(5.75 × 10–10 × 0.213)  = 1.11 × 10–5 molL-1  pH = –log[H3O+]  pH = 4.96 | * One step correct. | * TWO steps correct. | * All correct. |
| (c)  (d) | Since B is half way to the equivalence point, [NH4 +] = [NH3].  OR p*K*a = pH + log [acid]÷[c.base]  *so Ka = [H3O+]*  therefore p*K*a = pH.  *The solution at the equivalence point is NH4Cl.*  NH4+ solution is acidic since,  NH4+ + H2O  NH3 + H3O+ | * EITHER   *Ka* expression rearranged  OR  [NH4 +] = [NH3] at B  OR  pH = p*K*a   * NH4+ is acidic.   OR  Correct equation. | * [NH4+] = [NH3]   AND  pH = p*K*a   * NH4+ is acidic  AND  equation in either symbols or words. | * Complete explanation. |

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| **2015** | **Evidence** | **Achievement** | **Achievement with Merit** | **Achievement with Excellence** |
| (a)(i) | Na+, F–, H2O, HF, OH–, H3O+. | * Three species listed. |  |  |
| (ii) | A weak base, F–, is present at the equivalence point:  F– + H2O ⇌ 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) | pH = p*K*a + log [F–] / [HF]  = 3.17 + log 0.5  = 2.87  [H3O+] = 2 × 10–3.17 = 1.35 × 10–3 mol L–1  pH = –log (1.35 × 10–3) = 2.87.  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 p*K*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) | [OH–] = 0.258 × 4 / 44  = 0.02345 mol L–1 pOH = –log (0.02345)  = 1.63  pH = 14 – pOH  = 12.4 | * Correct *n*(NaOH).   OR  One correct step. | * Correct process but one error in calculation. | * Correct answer. |
| (b) | Since CH3COOH has a higher p*K*a, it is a weaker acid than HF. Therefore its conjugate base, CH3COO–, will be a stronger base than F–. This means [OH–] will be higher at the equivalence point for the CH3COOH vs NaOH titration, so the equivalence point pH will be higher. | * States CH3COOH is a weaker acid than HF.   OR  States equivalence point is higher with some evidence. | * Links higher equivalence point for CH3COOH to the strength of the weak acid, its p*K*a, and the relative strength of its conjugate base. | * Contrasts the equivalence point for both titrations. |

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| **2014** | **Evidence** | **Achievement** | **Achievement with Merit** | **Achievement with Excellence** |
| (a) | At point **A**, [CH3NH2] ≈ [CH3NH3+]. So the solution has buffering properties in the proximity of point **A**. When HBr is added, the H3O+ is consumed:  H3O+ + CH3NH2 → CH3NH3+ + H2O  Since the H3O+ is removed from the solution (neutralised), the pH does not change significantly. | * Recognises near point **A** solution is a buffer /   [CH3NH2] ≈ [CH3NH3+].   * Identifies H3O+ or HBr is neutralised / removed by CH3NH2. * Calculates [OH–] / [H3O+] / *K*b * Uses suitable process with more than one error.   OR  Rearranges *K*b / *K*a expression so [CH3NH2] is the subject. | * Correct equation linked to neutralisation / absorption of H3O+. * Correct method but an error in the calculation. | * Correct answer with a clear method. |
| (b) | [H3O+] = 10–11.8 = 1.58 x 10-12  *K*a =  =  [OH-]  2.29 × 10–11 =  [CH3NH2] =  = 0.0912 mol L–1  **OR**  [OH–] =  = 6.31 × 10–3 mol L–1  *K*b =  4.37 × 10–4 =  [CH3NH2] =  [CH3NH2] = 0.0912 mol L–1 |
| (c)(i)  (ii) | CH3NH3+, Br–, CH3NH2, H3O+ | 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 CH3NH2 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 CH3NH2 solution and the solution at point **B**, including a consideration of the differing concentrations of each solution. |
| **At the start**, before addition of HBr there is a solution of weak base (CH3NH2) which only partially reacts with water to produce a relatively low concentration of ions.  As a result, the initial CH3NH2 solution will be a poor electrical conductor.  **CH3NH2 + H2O ⇌ CH3NH3+ + OH-**  Therefore species present are CH3NH2 > OH– ≥ CH3NH3+ > H3O+  **At point B,** there is a solution of the salt CH3NH3Br present which is dissociated completely into ions. Therefore there is a relatively high concentration of ions (CH3NH3+ and Br–) present in the solution, so it will be a good electrical conductor / electrolyte.  **CH3NH3Br → CH3NH3+ + Br–**  CH3NH3+ reacts with water according to the equation  **CH3NH3+ + H2O ⇌ CH3NH2 + H3O+**  Species present are Br– > CH3NH3+ > H3O+ ≥ CH3NH2 > (OH–) |

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| **2013** | **Evidence** | **Achievement** | **Achievement with Merit** | **Achievement with Excellence** |
| (a) |  | * Correct process. | * Correct pH. | (a) & (b) correct. |
| (b) | Halfway to equivalence point, half of the ethanoic acid has been used up. There are now equimolar quantities of ethanoic acid and sodium ethanoate.    According to the equation when [CH3COOH] = [CH3COO–]  then *K*a = [H3O+]  So p*K*a = pH. | * Recognises that there are equimolar quantities of ethanoic acid and sodium ethanoate. | * Relates equation correctly to explanation. |  |
| (c)(i) | NaOH(*aq*) + CH3COOH(*aq*) → NaCH3COO(*aq*) + H2O(*l*) (1)  [CH3COO–] increases as it is formed in reaction (1).  [Na+] increases as NaOH is added (1).  [CH3COOH] decreases as it reacts with NaOH (1).  [H3O+] decreases because [CH3COO–] / [CH3COOH] increases and *K*a is a constant.  [OH–] increases because [H3O+] decreases and [H3O+] [OH–] is constant. | * Correct equation *minor error.* * Correct statement relating to change in concentration of 1 species. | Correct equation and correctly describes the change in concentration of 2 species. | * Correct equation.   AND  Correctly describes the change in concentration of the 4 species. |
| (c)(ii) | *n*(CH3COOH at start) = 0.0896 × 20 ×10–3  = 1.79 × 10–3 mol  *n*(NaOH added) = 0.1 × 5 × 10–3  = 5 × 10–4 mol  After 5 mL NaOH added:  *n*(CH3COOH) = 1.29 × 10–3 mol  *n*(CH3COO–) = 5 × 10–4 mol  [CH3COOH] = 0.0516 mol L–1  [CH3COO–] = 0.0200 mol L–1  [H3O+] = 4.48 × 10–5 mol L–1  pH = 4.35  Candidates will not be penalised for not calculating concentrations. | * Correct n for CH3COOH OR NaOH at the start. | * Correct process to identify either of the species after 5 mL has been added (mol or mol L–1). | * Correct answer. |

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| **2012** | **Evidence** | **Achievement** | **Merit** | **Excellence** |
| **(a) (i)**  **(ii)** | n (NaOH) = 0.180 mol L–1 × 0.04 L = 0.0072 mol  1NaOH : 1HCOOH  c (HCOOH) = 0.0072 mol / 0.025 L = 0.288 mol L–1  HCOOH + H2O ⇌ HCOO– + H3O+  *K*a = [HCOO–][H3O+] / [HCOOH]  [H3O+] = √(0.288 mol L–1 × 1.82 × 10–4) = 7.24 × 10–3 mol L–1  pH = 2.14  Half way to the equivalence point (20 mL), the NaOH has reacted with half the HCOOH, so [HCOOH] = [HCOO–].  *K*a = [HCOO–] [H3O+] / [HCOOH]  Since [HCOOH] = [HCOO–], substitution into *K*a gives:  *K*a = [H3O+]  Taking –log of each side:  p*K*a = pH  So the pH half way to the equivalence point = p*K*a of HCOOH = 3.74  According to the titration curve, the pH at the equivalence point is  approximately 8.4. Indicators change colour at their p*K*a +/–1. So  bromocresol green changes colour over a pH range of 3.7-5.7 and alizarin yellow will change colour over a pH range of 10.0 – 12.0. This means bromocresol green would therefore change colour before the equivalence point and alizarin yellow would change colour after the equivalence point.  Both of these indicators would therefore be unsuitable for this titration.  Cresol red will change colour over a pH range of 7.3 – 9.3. This includes the pH at the equivalence point. Cresol red would therefore be a suitable indicator to detect the endpoint since it changes colour at the equivalence point. | Any TWO of:  • Shows [HCOOH] = 0.288 molL–1  • Method to find pH correct with  minor arithmetic error.  • EITHER:  States [HCOOH] = [HCOO–]  OR  pH = p*K*a  EITHER:  Identifies correct or incorrect  indicator(s) with limited reasoning.  OR  Identifies that the endpoint needs to be close to the equivalence point. | Any TWO of:  • Correct pH.  States [HCOOH] = [HCOO–] so pH = p*K*a when half-way to the equivalence point / in buffer zone.  Identifies the correct indicator and an incorrect indicator, including  consequences of wrong choice, but gaps in discussion (i.e. have talked  in general terms about an  incorrect choice). | 1 answer to merit level  AND  Discussion of indicators links the  equivalence point pH to the effective p*K*a over which each  indicator will change colour (range may be for either the indicator colour change occurring over a range  OR  for it’s p*K*a falling within the range of the vertical part of graph). |

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| **2011** | **Evidence** | **Achievement** | **Merit** | **Excellence** |
|  | **A**  At point A, there is an equi-molar mixture of HEt and Et–. On addition of OH– ions, the acid part of the buffer neutralises the OH– ions, by donating a proton. The acid reacts with the base:  HEt + OH– → Et– + H2O  On addition of H3O+, the ethanoate will accept a proton from the hydronium ion:  Et– + H3O+ → HEt + H2O  Candidate may discuss equilibrium shift.  p*K*a = pH = 4.76 (accept 4.5 – 4.9)  **B**  At the equivalence point all the HEt has been  neutralised by NaOH.  HEt + NaOH → EtNa + H2O  The Et– reacts further to a small extent with water.  Et– + H2O ⇌ HEt + OH–  Thus the pH of the equivalence point is above 7 due to  presence of OH–. | ONE of:  • Recognises that at point A there is a buffer solution.  • States that equimolar amounts of  acid / base conjugate are present at A.  • States that pH will not change when small amounts of acid or base are added.  • Correct p*K*a / Ka  AND  ONE of:  • Recognises that all the HEt has been  used up at B.  • That the pH of equivalence point is  greater than 7. (must have clearly  indicated that point B is the  equivalence point) | Describes how a buffer works  (for when both acid AND base  are added) by:  EITHER  • Giving equations for the specific buffer  OR  • Writing about how a buffer works in general terms  OR  • Links that due to equimolar  HEt and Et– thus pKa = pH  AND  • Recognises that none of the  original HEt remains as ithas  all reacted with NaOH  OR  • That the pH of equivalence  point is greater than 7 with a  valid reason. | Shows recognition of equimolar  HEt and Et– thus p*K*a = pH and discusses how the buffer solution works and links to equations.  AND  Uses two equations to explain  why the pH is above 7. (One  equation may be implied in the  candidate’s written answer.) |

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| **2010** | **Evidence** | **Achieve** | **Merit** | **Excellence** |
| (a) | Since B is half way to the equivalence point,  [NH4+] = [NH3]  *K*a =  so *K*a = [H+] p*K*a = pH.  Note: [NH4+] = [NH3] must include [ ] or the word **concentration**. | Any TWO of:   * EITHER   pH = p*K*a  OR  [NH4+] = [NH3]  at B. | Any TWO of:   * Relates pH half way to equivalence point to species present to give  pH = p*K*a |  |
| (b) | p*K*b(NH3) = 4.76 *K*b = 1.74 × 10–5  NH3 + H2O sm eq arrow NH4+ + OH–  *K*b =  [OH–] =  = 1.67 × 10–3 mol L–1  pOH = 2.8 (accept 2.77) **pH = 11.2** (accept 11.22)  OR  *K*a = 5.75 × 10–10 | * EITHER   Rearranges *K*a / *K*b expression so [H3O+] / [OH–] is subject.  OR  *K*b / p*K*b calculated  OR  Recognition that *K*w=  [H3O+] [OH–]  OR  Recognition that  14 = pH + pOH. | * EITHER   Correct method with error in calculation.  OR  Calculates [H3O+] / [OH–]. | Calculates pH at point A. |
| (c) | Solution at equivalence point is NH4Cl. This solution is weakly acidic since  NH4+ + H2O sm eq arrow NH3 + H3O+.  so pH < 7. | * EITHER   pH < 7 as all of initial NH3 has reacted / **equal amount** of HCl has been added.  OR  ‘e’ answer without statement that pH < 7. | * Correct pH related to species present. | AND  pH < 7 as all of initial NH3 has reacted to form NH4+ / NH4Cl AND that NH4+ is a weak acid AND equation for NH4+ dissociation. |

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| **2009** | **Evidence** | **Achievement** | **Merit** | **Excellence** |
| (a) | NaOH → Na+ + OH–  CH3NH2 + H2O ⇌ CH3NH3+ + OH–  CH3NH3Cl → CH3NH3+  + Cl–  CH3NH3+  + H2O ⇌ CH3NH2 + H3O+ | * Any 2 equations correct. * Recognises reasons for pH variation are due to differences in H3O+ or OH– concentrations   **OR**   * Recognises conductivity is related to the concentration of ions in solution. | * Equations correct for all solutions. * Recognises reasons for variations in pH and conductivity and makes one valid comparison.   **OR**   * Difference in conductivity   correctly discussed relating to the specific solutions  **OR**   * Difference in pH correctly discussed relating to the specific solutions. | * Discussion addresses variation in BOTH pH and conductance using correct reasons. |
| (b) | NaOH is a strong base and completely dissociates into its ions. Strong base, so produces more OH– ions in solution than CH3NH2. pH > 7 and greater than for CH3NH2.  CH3NH2 is a weak base, so only some ions react with water to produce OH– ions. pH > 7.  CH3NH3Cl is an acidic salt that completely dissociates into its ions producing CH3NH3+.  CH3NH3+  is a weak acid, so partially dissociates in water to produce H3O+ ions. So pH < 7.  Conductivity relates to the number of ions in a solution.  NaOH and CH3NH3Cl completely dissociate, so produce a large number of ions, so high conductivity.  CH3NH2 has an equilibrium reaction with water, so only a few ions are produced, so low conductivity. |

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| 2008 | **Evidence** | **Achievement** | **Achievement with Merit** | **Achievement with Excellence** |
| (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.  14  12  10  8  6  4  2  0   |  |  |  |  |  |  |  | | --- | --- | --- | --- | --- | --- | --- | |  |  |  |  |  |  |  | |  |  |  |  |  |  |  | | pH |  |  |  |  |  |  | |  |  |  |  |  |  |  | |  |  |  |  |  |  |  | |  |  |  |  |  |  |  | |
| (b) | HN3(*aq*) + H2O H3O+(*aq*) + N3–(*aq*)  *K*a(CH3COOH) = [H3O+] [N3–]  [HN3]  pH = 2.6 so [H3O+(*aq*)] = **2.51 × 10–3 mol L–1** = [N3–(*aq*)]  p*K*a(HN3) = 4.72, *K*a(HN3) = **1.91 × 10–5**  *K*a(HN3) = 1.91 × 10–5  = (2.51 × 10–3)2  *x*  ***x* = 0.331 ie [HN3 (*aq*)] = 0.331 mol L–1** (Allow 0.330 and 0.33) | [H3O+] determined from pH  OR  *K*a determined from p*K*a.  (As long as solution is identified.) | Concentration of acid solution correctly determined.  (Must have correct unit) |  |

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