Chemistry 3.6 AS 91392

Demonstrate understanding of equilibrium principles in aqueous systems

WORKBOOK

Working to Excellence & NCEA Questions

CONTENTS (including NCEA Questions under each topic)

GZ Science Resources

Writing Excellence answers to **Solubility of sparingly soluble salts** questions

Solubility of sparingly soluble salts QUESTION	
Question: Silver carbonate, Ag_2CO_3, is a sparingly soluble salt. $K_s(Ag_2CO_3) = 8.10 \times 10^{-12}$ at 25ºC $M(Ag_2CO_3) = 276$ g mol^{-1} (a) Write the solubility product expression, K_s, for silver carbonate (Ag_2CO_3). (b) Calculate the mass of Ag_2CO_3 that will dissolve in 50 mL of water to make a saturated solution at 25ºC.	
ANSWER	
1. write the equation for the dissociation of salt	
2. Write the solubility product expression, K_s, for the salt	
3. calculate the solubility, s 2:1 salt Let s = solubility $K_s = 4s^3$ *3sgf and units*	
4. calculate number of moles $n = c \times v$ *3sgf and units*	
5. calculate mass of salt $m = n \times M$ *3sgf and units*	

NOTE: The white column is how your answer would appear on your test paper so make sure you **write out complete sentences**. The grey area is just to help you structure your answer and would not appear in the question.

Past NCEA questions **Solubility of sparingly soluble salts**

2013: 2a. In an experiment, a saturated solution was made by dissolving 1.44×10^{-3} g of Ag_2CrO_4 in water, and making it up to a volume of 50.0 mL.
M (Ag_2CrO_4) = 332 g mol^{-1}
(a) Write the K_s expression for $Ag_2CrO_{4(s)}$.
(b) i. Calculate the solubility of $Ag_2CrO_{4(s)}$, and hence give the $[Ag^+]$ and $[CrO_4^{2-}]$ in the solution.
(b) ii. Determine the $K_s(Ag_2CrO_4)$.

2014: 2a. A flask contains a saturated solution of $PbCl_2$ in the presence of undissolved $PbCl_2$.
(i) Write the equation for the dissolving equilibrium in a saturated solution of $PbCl_2$.
(ii) Write the expression for $K_s(PbCl_2)$.
(iii) Calculate the solubility (in mol L^{-1}) of lead(II) chloride in water at 25°C, and give the $[Pb^{2+}]$ and $[Cl^-]$ in the solution.
$K_s(PbCl_2) = 1.70 \times 10^{-5}$ at 25°C

2015: 2a. Sufficient calcium carbonate, $CaCO_{3(s)}$, is dissolved in water to make a saturated solution.
(i) Write the equation for the equilibrium occurring in a saturated solution of $CaCO_3$.
(ii) Write the expression for $K_s(CaCO_3)$.
(iii) Calculate the solubility product of $CaCO_3$, $K_s(CaCO_3)$.
The solubility of $CaCO_3$ is 5.74×10^{-5} mol L^{-1}.

2016: 1a. Silver carbonate, Ag_2CO_3, is a sparingly soluble salt.
$K_s(Ag_2CO_3) = 8.10 \times 10^{-12}$ at 25ºC $M(Ag_2CO_3)$ = 276 g mol^{-1}
Write the solubility product expression, K_s, for silver carbonate (Ag_2CO_3).
1b. Silver carbonate, Ag_2CO_3, is a sparingly soluble salt.
$K_s(Ag_2CO_3) = 8.10 \times 10^{-12}$ at 25ºC $M(Ag_2CO_3)$ = 276 g mol^{-1}
Calculate the mass of Ag_2CO_3 that will dissolve in 50 mL of water to make a saturated solution at 25ºC.

2017: 1b. (i) Write the equation for the equilibrium occurring in a saturated solution of copper(II) hydroxide, $Cu(OH)_2$.
(ii) Write the expression for $K_s(Cu(OH)_2)$.
(iii) Calculate the solubility of $Cu(OH)_2$ in water at 25°C.
$K_s(Cu(OH)_2) = 4.80 \times 10^{-20}$

Writing Excellence answers to **Common Ion Effect** questions

Common Ion Effect QUESTION	
Question: Show, by calculation, that a precipitate of lead(II) hydroxide, $Pb(OH)_2$, will form when 25.0 mL of a sodium hydroxide solution, NaOH, at pH 12.6 is added to 25.0 mL of a 0.00421 mol L^{-1} lead(II) nitrate, $Pb(NO_3)_2$, solution. $K_s(Pb(OH)_2) = 8.00 \times 10^{-17}$ at 25°C	
ANSWER	
1. write the equation for the dissociation of salt	
2. Write the solubility product expression, *Q*, for the salt (K_s)	
3. calculate the solubility, s for the first ion after dilution $[Pb^{2+}] = \frac{c \times v}{\text{total } v}$ *3sgf and units*	
4. calculate the concentration of [OH-] from pH $[OH^-] = 10^{-(14-pH)}$ *3sgf and units*	
5. calculate the solubility, s for the second ion after dilution $[OH^-] = \frac{c \times v}{\text{total } v}$ *3sgf and units*	
6. Calculate Q from expression $Q = [ion1] \times [ion2]^2$ *3sgf (has no units)*	
7. compare Q and Ks and state whether a precipitate will form or not	

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Past NCEA questions **Common Ion Effect**

2014: 2b. A sample of seawater has a chloride ion concentration of 0.440 mol L^{-1}.

Determine whether a precipitate of lead(II) chloride will form when a 2.00 g sample of lead(II) nitrate is added to 500 mL of the seawater.
$K_s(PbCl_2) = 1.70 \times 10^{-5}$ $M(Pb(NO_3)_2) = 331$ g mol^{-1}

2015: 2c. Show, by calculation, that a precipitate of lead(II) hydroxide, $Pb(OH)_2$, will form when 25.0 mL of a sodium hydroxide solution, NaOH, at pH 12.6 is added to 25.0 mL of a 0.00421 mol L^{-1} lead(II) nitrate, $Pb(NO_3)_2$, solution.

$K_s(Pb(OH)_2) = 8.00 \times 10^{-17}$ at 25°C

2016: 1d. Show by calculation whether a precipitate of Ag_2CO_3 will form when 20.0 mL of 0.105 mol L^{-1} silver nitrate, $AgNO_3$, solution is added to 35.0 mL of a 0.221 mol L^{-1} sodium carbonate, Na_2CO_3, solution.

$K_s(Ag_2CO_3) = 8.10 \times 10^{-12}$ at 25ºC

2017: 1b (iii). 40.0 mL of 0.150 mol L–1 HBr solution was added to 25.0 mL of a saturated silver bromide, AgBr, solution.

Calculate the concentration of the silver ions, Ag^+, after the HBr solution has been added.

$K_s(AgBr) = 5.00 \times 10^{-13}$

Assume the concentration of Br^- in the original saturated solution of AgBr is insignificant.

Writing Excellence answers to **Solubility and Equilibrium** questions

Solubility and Equilibrium QUESTION	
Question: The solubility of zinc hydroxide, $Zn(OH)_2$, can be altered by changes in pH. Some changes in pH may lead to the formation of complex ions, such as the zincate ion, $[Zn(OH)_4]^{2-}$ Use equilibrium principles to explain why the solubility of zinc hydroxide increases when the pH is less than 4 or greater than 10.	
ANSWER	
1. write the **equation** for the dissociation of salt	
2. Explain that OH- ions are formed during dissociation	
3. write the **equation** for the reaction of H_3O^+ ions + OH^- ions when adding acid (due to pH being less than 4)	
4. link removal of OH^- ions (product) to **equilibrium** shifting AND change in solubility	
5. write the **equation** for the formation of the complex ion $[Zn(OH)_4]^{2-}$ with excess OH^- ions (due to pH being greater than 10)	
6. link removal of OH^- ions (product) to **equilibrium** shifting AND change in solubility	

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Past NCEA questions **Solubility and Equilibrium**

2013: 2c. In another experiment, 0.0100 g of Ag_2CrO_4 in beaker A was made up to a volume of 50.0 mL with water. In beaker B, 0.0100 g of Ag_2CrO_4 was made up to a volume of 50.0 mL with 0.100 mol L^{-1} ammonia solution. Compare and contrast the solubility of Ag_2CrO_4 in beaker A and beaker B.

2014: 2c. The solubility of zinc hydroxide, $Zn(OH)_2$, can be altered by changes in pH. Some changes in pH may lead to the formation of complex ions, such as the zincate ion, $[Zn(OH)_4]^{2-}$
Use equilibrium principles to explain why the solubility of zinc hydroxide increases when the pH is less than 4 or greater than 10.

2015: 2b. Some marine animals use calcium carbonate to form their shells. Increased acidification of the oceans poses a problem for the survival of these marine animals.

Explain why the solubility of $CaCO_3$ is higher in an acidic solution.

Use an equation to support your explanation.

2016: 1c. Explain how the solubility of Ag_2CO_3 will change if added to 50 mL of a 1.00 mol L^{-1} ammonia, NH_3, solution.

Support your answer with balanced equations.

2017: 1b. 40.0 mL of 0.150 mol L–1 HBr solution was added to 25.0 mL of a saturated silver bromide, AgBr, solution.

(i) Write an equation for the equilibrium occurring in a saturated solution of AgBr.

(ii) (ii) Explain the changes that occur to the concentrations of the species in the saturated solution of AgBr on the addition of the HBr solution.

2017: 2c. Explain why the solubility of $Cu(OH)_2$ increases when dilute hydrochloric acid is added.

Past NCEA questions dissociation equations

Candidates are expected to recognise common strong acids (HCl, HBr, HNO_3, H_2SO_4); strong bases (KOH, NaOH); weak acids (HF, CH_3COOH, and NH_4^+); weak bases (NH_3, CH_3NH_2, and CH_3COO^-). Less familiar weak acids and bases may be included in the context of appropriate resource information.

2014: 1a. When chlorine gas is added to water, the equation for the reaction is:
$Cl_{2(g)} + H_2O_{(l)} \rightleftharpoons HCl_{(aq)} + HOCl_{(aq)}$

(i) Write an equation for the reaction of the weak acid, hypochlorous acid, HOCl, with water

2015: 1a: Methylammonium chloride, CH_3NH_3Cl, dissolves in water to form a weakly acidic solution.
$Ka(CH_3NH_3^+) = 2.29 \times 10^{-11}$

(i) Write an equation to show CH_3NH_3Cl dissolving in water.

2016: 2a: (i) Ethanamine, $CH_3CH_2NH_2$, is a weak base.
$pK_a(CH_3CH_2NH_3^+) = 10.6$ $K_a(CH_3CH_2NH_3^+) = 2.51 \times 10^{-11}$

Write an equation to show the reaction of ethanamine with water

2017: 1a: Hydrogen fluoride, HF, and hydrogen bromide, HBr, both form acidic solutions when added to water.

(i) Write an equation for the reaction of each acid with water.
Hydrogen fluoride, HF, with water:
Hydrogen bromide, HBr, with water:

Writing Excellence answers to **Species in Solution** questions

Concentration of Species QUESTION	
Question: Ethyl ammonium chloride, $CH_3CH_2NH_3Cl$, is a weak acid that will also react with water. List all the species present in a solution of $CH_3CH_2NH_3Cl$, in order of decreasing concentration. Do not include water. Justify the order you have given. Include equations, where necessary.	
ANSWER	
1. write the **equation** for the dissociation of salt	
2. link to complete dissociation AND formation of an **(spectator) ion** that does not react further so will be in greatest concentration	
3. write the **equation** for the weak acid (formed from equation above) in water	
4. link to partial dissociation due to being a **weak acid** AND most will remain so will be next in concentration	
5. Explain **H_30^+ ions** are formed during reaction in same quantity as conjugate PLUS small contribution from water AND so will be next in concentration	
6. Explain **conjugate base** are formed during reaction in same quantity as H_30^+ AND so will be next in concentration (but both H_30^+ ions and conjugate will be at smaller concentration to acid as only weak acid)	
7. Finally Explain **OH- ions** present in small amounts from water dissociation only AND so will be last in concentration	
8. list **species in order**	

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Past NCEA questions **Species in Solution**

2013: 1a. 1 mol of each of the following substances was placed in separate flasks, and water was added to these flasks to give a total volume of 1 L for each solution. In the box below, rank these solutions in order of increasing pH. Justify your choice and include equations where appropriate.

CH_3NH_3Cl CH_3NH_2 HCl

2014: 1a. When chlorine gas is added to water, the equation for the reaction is:

$Cl_{2(g)} + H_2O_{(l)} \rightleftharpoons HCl_{(aq)} + HOCl_{(aq)}$

(ii) List all the species present when HOCl reacts with water, in order of decreasing concentration. Justify your order.

2015: 1a (iii). List all the species present in an aqueous solution of CH_3NH_3Cl, in order of decreasing concentration.

Do not include water.

2016: 2c. Ethyl ammonium chloride, $CH_3CH_2NH_3Cl$, is a weak acid that will also react with water.

List all the species present in a solution of $CH_3CH_2NH_3Cl$, in order of decreasing concentration.

Do not include water.

Justify the order you have given.

Include equations, where necessary.

2017: 3c The equivalence point pH for the titration of ethanoic acid with sodium hydroxide is 8.79.

(i) Identify the chemical species present at the equivalence point, other than water.

Writing Excellence answers to **Conductivity and Ions** questions

Conductivity and Ions QUESTION

Question: The table shows the pH and electrical conductivity of three solutions. The concentrations of the solutions are the same. Compare and contrast the pH and electrical conductivity of these three solutions. Include appropriate equations in your answer.

Solution	NaOH	CH_3NH_2	CH_3COONa
pH	13.2	11.9	8.98
Electrical conductivity	good	poor	good

ANSWER

1. Identify each solution as either being a **weak or strong acid or base (or salt)** linked to the **pH** (and presence of ions)	
2. State requirements for **conductivity**	
3. **Solution NaOH** (pH 13.2) Write equation **AND** link ions formed to conductivity and level of dissociation	
4. pH **Solution NaOH** (pH 13.2) Link amounts of H_3O_+ / OH^- ions to pH	
5. **Solution CH_3NH_2** (pH 11.9) Write equation **AND** link ions formed to conductivity and level of dissociation	
6. pH **Solution CH_3NH_2** (pH 11.9) Link amounts of H_3O_+ / OH^- ions to pH (compared to previous solution)	
7. **Solution CH_3COONa** (pH 8.98) Equation 1. [salt dissociation] Write equation **AND** link ions formed to conductivity and level of dissociation	
8. **Solution CH_3COONa** (pH 8.98) Equation 2.[acid reaction] Write equation **AND** link ions formed to conductivity and level of dissociation	
9. pH **Solution H_3COONa** (pH 8.98) Link amounts of H_3O_+ / OH^- ions to pH (compared to previous solution)	

Past NCEA questions **Conductivity and Ions**

2013: 1b. The conductivity of the 1 mol L–1 solutions formed in (a) can be measured.

CH_3NH_3Cl CH_3NH_2 HCl

Rank these solutions in order of decreasing conductivity. Compare and contrast the conductivity of each of the 1 mol L^{-1} solutions, with reference to species in solution.

2015: 1b. The table shows the pH and electrical conductivity of three solutions. The concentrations of the solutions are the same. Compare and contrast the pH and electrical conductivity of these three solutions. Include appropriate equations in your answer.

Solution	**NaOH**	**CH_3NH_2**	**CH_3COONa**
pH	13.2	11.9	8.98
Electrical conductivity	good	poor	good

2017: 1a. (ii) : Compare and contrast the electrical conductivity of 0.150 mol L–1 solutions of hydrofluoric acid, HF, and hydrobromic acid, HBr.

In your answer, you should:

• include the requirements for a solution to conduct electricity

• identify the species present AND their relative concentrations.

No calculations are necessary.

Writing Excellence answers to **pH Calculations** questions

pH Calculations QUESTION 1. (4 steps excellence)	
Question: Calculate the pH of a 0.109 mol L^{-1} solution of ethanamine. $pK_a(CH_3CH_2NH_3^+) = 10.6$ $K_w = 1.00\times10^{-14}$	
ANSWER	
1. determine if the solution is acid or base (will it accept or donate H^+) – strong or weak And write down all available information	
2. convert pK_a to K_a $K_a = 10^{-pKa}$	
3. calculate $[H_3O^+]$ $[H_3O^+] = \sqrt{\frac{Ka \times Kw}{[base]}}$ *3sgf and units*	
4. calculate pH pH = -log $[H_3O^+]$ *3sgf* *Double check answer against expected pH for your solution*	

pH Calculations QUESTION 2. (3 steps Merit)	
Question: Calculate the pH of 0.0152 mol L^{-1} CH_3NH_3Cl solution. $K_a(CH_3NH_3^+) = 2.29 \times 10^{-11}$	
ANSWER	
1. determine if the solution is acid or base (will it accept or donate H^+) – strong or weak And write down all available information	
2. calculate $[H_3O^+]$ $[H_3O^+] = \sqrt{K_a \times c(HA)}$ *3sgf and units*	
3. calculate pH pH = -log $[H_3O^+]$ *3sgf* *Double check answer against expected pH for your solution*	

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Past NCEA questions pH Calculations

$K_a = 10^{-Pka}$

$[H_3O^+] = \sqrt{K_a \times c(HA)}$

$pH = -\log[H_3O^+]$

$K_a = 10^{-Pka}$

$[H_3O^+] = \sqrt{(K_a \times K_w \div [B])}$

$pH = -\log[H_3O^+]$

2014: 1a. Hypochlorous acid has a p*K*a of 7.53. Another weak acid, hydrofluoric acid, HF, has a p*K*a of 3.17.

A 0.100 mol L–1 solution of each acid was prepared by dissolving it in water.
Compare the pHs of these two solutions.
No calculations are necessary.

2015: 1a: (iv) Calculate the pH of 0.0152 mol L^{-1} CH_3NH_3Cl solution.

$Ka(CH_3NH_3^+) = 2.29 \times 10^{-11}$

2016: 2b: Calculate the pH of a 0.109 mol L^{-1} solution of ethanamine.

$pK_a(CH_3CH_2NH_3^+) = 10.6$

2017: 2a: Ammonia, NH_3, is a weak base.

$pK_a(NH_4^+) = 9.24$ $K_a(NH_4^+) = 5.75 \times 10^{-10}$
(i) Calculate the pH of a 0.105 mol L^{-1} NH_3 solution.

Writing Excellence answers to **Titration Curve – Start pH** questions

Titration Curve – Start pH QUESTION	
Question: A titration was carried out by adding hydrobromic acid, HBr, to 20.0 mL of aqueous methylamine, CH_3NH_2, solution. The equation for the reaction is: $CH_3NH_2 + HBr \rightarrow CH_3NH_3^+ + Br^-$ $K_a(CH_3NH_3^+) = 2.29 \times 10^{-11}$ $K_w = 1.00 \times 10^{-14}$ The aqueous methylamine, CH_3NH_2, solution has a pH of 11.8 before any HBr is added. Show by calculation that the concentration of this solution is 0.0912 mol L^{-1}.	
ANSWER	
1. determine if starting solution is acid or base (will it accept or donate H^+) – strong or weak And write down all available information	
2. calculate $[H_3O^+]$ $[H_3O^+] = 10^{-pH}$ *3sgf and units*	
3. write out K_a expression $K_a = \frac{[base][H_3O^+]}{[conj\ acid]}$ And then $K_a = \frac{[base][H_3O^+]}{[OH^-]}$	
4. rearrange to calculate $[CH_3NH_2]$ $[CH_3NH_2] = \frac{K_a \times K_w}{[H_3O^+]^2}$ *Assumptions: $[conj] = [H_3O^+]$* *$[OH^-] = K_w / [H_3O^+]$* *3sgf and units*	

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Writing Excellence answers to **Titration Curve – After the Start pH** questions

Titration Curve – after the Start pH QUESTION	
Question: : 20.0 mL of 0.0896 mol L^{-1} ethanoic acid is titrated with 0.100 mol L^{-1} sodium hydroxide. p*K*a (CH_3COOH) = 4.76 Calculate the pH of the titration mixture after 5.00 mL of NaOH has been added. $K_w = 1 \times 10^{-14}$	
ANSWER	
1. determine if starting solution is acid or base (will it accept or donate H^+) – strong or weak And write down all available information	
2. Write down neutralisation equation	
3. calculate *n*(CH_3COOH at start) n = cv *3sgf and units*	
4. calculate *n*(NaOH) and therefore n(CH_3COO-) n = cv *assume* *n*(NaOH) = *n*(CH_3COO-) *3sgf and units*	
5. calculate *n*(CH_3COOH) After 5 mL NaOH added: (total 25mL) =(n(CH_3COOH – n(CH_3COO-) after 5mL) *3sgf and units*	
6. calculate c(CH_3COOH) c = n/v (remember v = 25mL) *3sgf and units*	
7. calculate c(CH_3COO^-) c = n/v (remember v = 25mL) *3sgf and units*	
8. Calculate pH pK_a = 4.76 $pH = pK_a + \log \frac{[CH_3COO^-]}{[CH_3COOH]}$ *3sgf* *Check pH against estimate on curve*	

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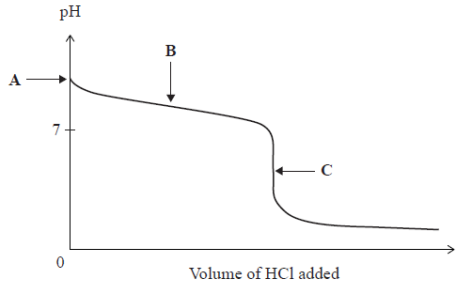
Writing Excellence answers to **Titration Curve – Equivalence Point pH** questions

Titration Curve – Equivalence Point pH QUESTION
Question: 20.00 mL of 0.320 mol L^{-1} ammonia, NH_3, is titrated with 0.640 mol L^{-1} hydrochloric acid, HCl. p$K_a(NH_4^+)$ = 9.24 Show, by calculation, that the pH at the equivalence point (point C) is 4.96. $K_w = 1 \times 10^{-14}$ 

ANSWER	
1. determine if equivalence point is greater or less than 7 (from curve or strong base/weak acid strong acid/weak base And write down all available information	
2. Write down neutralisation equation	
3. calculate *n*(Base) to neutralise (and reach equivalence point and therefore n(Acid) from 1:1 equation) n = cv *also assume* $n(NH_3) = n(NH_4^+)$ *3sgf and units*	
4. calculate *v*(Acid) to neutralise ($n(NH_3)$ = n(HCl) from 1:1 equation) v=n/c *3sgf and units*	
5. calculate $c(B^+)$ c = n/total v *also assume* $n(B) = n(B^+)$ see step 3. B = NH_3 B+ = NH_4^+ total v = start volume base + v acid added *3sgf and units*	
6. calculate $[H_3O^+]$ $K_a = 10^{-pKa}$ $[H_3O^+] = \sqrt{Ka \times c(B^+)}$ *3sgf and units* $B^+ = HA$	
7. Calculate pH pH = -log $[H_3O^+]$ *3sgf* *Check pH against estimate on curve*	

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Past NCEA questions **Titration Calculations**

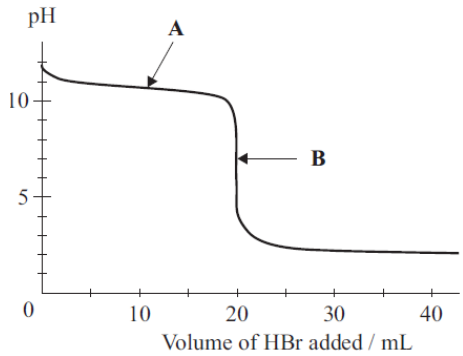
2013:3a. 20.0 mL of 0.0896 mol L^{-1} ethanoic acid is titrated with 0.100 mol L^{-1} sodium hydroxide.
p*K*a (CH_3COOH) = 4.76 Calculate the pH of the ethanoic acid before any NaOH is added.

2013:3b. Halfway to the equivalence point of the titration, the pH = p*K*a okf the ethanoic acid.
Discuss the reason for this.

2013:3c. (i) Discuss the change in the concentration of species in solution, as the first 5.00 mL of NaOH is added to the 20.0 mL of ethanoic acid. Your answer should include chemical equations. No calculations are required.

(ii) Calculate the pH of the titration mixture after 5.00 mL of NaOH has been added.

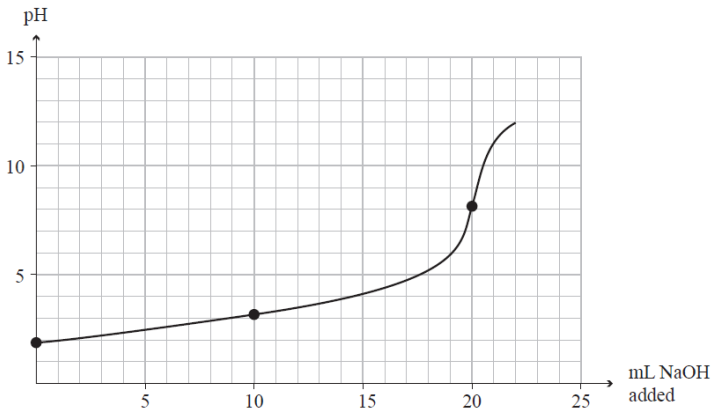
2014:3. A titration was carried out by adding hydrobromic acid, HBr, to 20.0 mL of aqueous methylamine, CH_3NH_2, solution.
The equation for the reaction is:
$CH_3NH_2 + HBr \rightarrow CH_3NH_3^+ + Br^-$ $K_a(CH_3NH_3^+) = 2.29 \times 10^{-11}$
(a) Explain why the pH does not change significantly between the addition of 5 to 15 mL of HBr (around point A on the curve).
(b) The aqueous methylamine, CH_3NH_2, solution has a pH of 11.8 before any HBr is added.
Show by calculation that the concentration of this solution is 0.0912 mol L^{-1}.



2014:3. (c) (i) Write the formulae of the four chemical species, apart from water and OH–, that are present at the point marked B on the curve (above).
(ii) Compare and contrast the solution at point B with the initial aqueous methylamine solution.
In your answer you should include:
- a comparison of species present AND their relative concentrations
- a comparison of electrical conductivity linked to the relevant species present in each solution
- equations to support your answer

2015: 3a: 20.0 mL of 0.258 mol L^{-1} hydrofluoric acid, HF, solution is titrated with a sodium hydroxide, NaOH, solution.

The equation for the reaction is:
$HF + NaOH \rightarrow NaF + H_2O$
pK_a(HF) = 3.17
(i) Identify the species in solution at the equivalence point.



(ii) Explain why the pH at the equivalence point is greater than 7. Include an equation in your answer.

2015: 3a: 20.0 mL of 0.258 mol L^{-1} hydrofluoric acid, HF, solution is titrated with a sodium hydroxide, NaOH, solution. $HF + NaOH \rightarrow NaF + H_2O$ p*K*a(HF) = 3.17

(iii) After a certain volume of NaOH solution has been added, the concentration of HF in the solution will be twice that of the F–. Calculate the pH of this solution, and evaluate its ability to function as a buffer.

(iv) Determine by calculation, the pH of the solution after 24.0 mL of 0.258 mol L^{-1} NaOH solution has been added. .

3b: In a second titration, a 0.258 mol L^{-1} ethanoic acid, CH_3COOH, solution was titrated with the NaOH solution. Contrast the expected pH at the equivalence point with the HF titration.

p*K*a(CH_3COOH) = 4.76 No calculations are necessary.

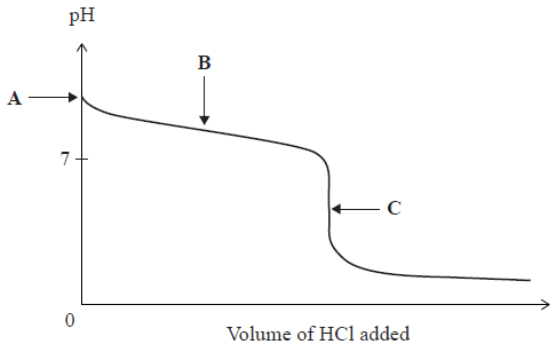
2016: 3a: 20.00 mL of 0.320 mol L^{-1} ammonia, NH_3, is titrated with 0.640 mol L^{-1} hydrochloric acid, HCl.

The equation for this reaction is:

$NH_3 + HCl \rightarrow NH_4^+ + Cl^-$ p$K_a(NH_4^+)$ = 9.24

The curve for this titration is given below.

Explain why the pH at the equivalence point (point C) is not 7.



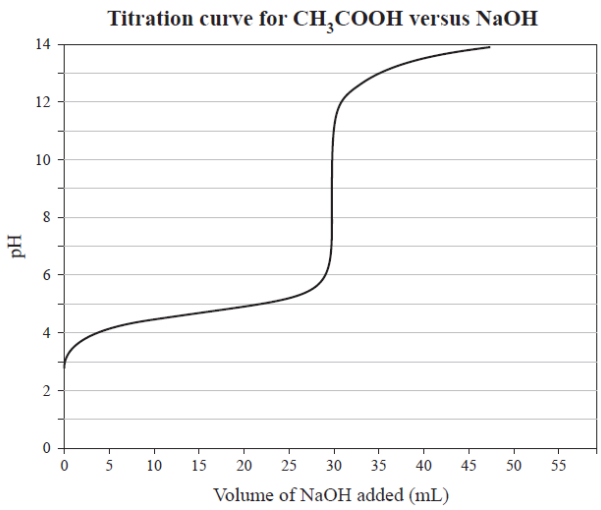
2016: 3b. Explain, in terms of the species present, why the pH at B (half way to the equivalence point) is 9.24.

3c. Show, by calculation, that the pH at the equivalence point (point C) is 4.96.

3d. Explain, in terms of the species present, why the pH of the solution at point C is 4.96.

2017: 3a: A titration was carried out by adding 0.112 mol L^{-1} sodium hydroxide solution, $NaOH_{(aq)}$, to 20.0 mL of ethanoic acid solution, $CH_3COOH_{(aq)}$. The equation for the reaction is:

$CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow CH_3COONa_{(aq)} + H_2O_{(l)}$ $K_a(CH_3COOH) = 1.74 \times 10^{-5}$



(a) With reference to the titration curve above, put a tick next to the indicator most suited to identify the equivalence point.

Indicator	pK_a	Tick ONE box below
Methyl yellow	3.1	
Bromocresol purple	6.3	
Phenolphthalein	9.6	

(b) (i) The ethanoic acid solution, $CH_3COOH_{(aq)}$, has a pH of 2.77 before any NaOH is added. Show by calculation that the concentration of the CH_3COOH solution is 0.166 mol L^{-1}.

(ii) Calculate the pH of the solution in the flask after 10.0 mL of 0.112 mol L^{-1} NaOH has been added to 20.0 mL of ethanoic acid solution, $CH_3COOH_{(aq)}$.

2017: 3c: The equivalence point pH for the titration of ethanoic acid with sodium hydroxide is 8.79.

(i) Identify the chemical species present at the equivalence point, other than water.

(ii) In a second titration, a 0.166 mol L^{-1} methanoic acid solution, $HCOOH_{(aq)}$, is titrated with the NaOH solution. The equivalence point pH for this titration is 8.28.

The equivalence point pH for the CH_3COOH titration is 8.79.

Compare and contrast the pH values at the equivalence point for both titrations.

$K_a(HCOOH) = 1.82 \times 10^{-4}$ $K_a(CH_3COOH) = 1.74 \times 10^{-5}$ *No calculations are necessary.*

Writing Excellence answers to **Buffer pH Calculation** questions

Buffer pH Calculation QUESTION	
Question: The following two solutions from part (a) are mixed to form a buffer solution: 20.0 mL of 1 mol L^{-1} CH_3NH_3Cl and 30.0 mL of 1 mol L^{-1} CH_3NH_2 Calculate the pH of the resultant buffer solution. p*K*a ($CH_3NH_3^+$) = 10.64 $K_w = 1 \times 10^{-14}$	
ANSWER	
1. Write out Ka expression	
2. rearrange expression to calculate $[H_3O^+]$	
3. calculate $[CH_3NH_2]$ $[CH_3NH_2] = \frac{v(CH_3NH_2) \times c(CH_3NH_2)}{\text{total volume}}$ *3sgf and units*	
4. calculate $[CH_3NH_3^+]$ $[CH_3NH_3^+] = \frac{v(CH_3NH_3^+) \times c(CH_3NH_3^+)}{\text{total volume}}$ *3sgf and units*	
5. calculate *pH* $pH = pKa + \log \frac{[A^-]}{[HA]}$ *3sgf*	

NOTE: The white column is how your answer would appear on your test paper so make sure you **write out complete sentences**. The grey area is just to help you structure your answer and would not appear in the question.

Past NCEA questions **Buffer Calculations**

2013:1c: (i) The following two solutions from part (a) are mixed to form a buffer solution:
20.0 mL of 1 mol L^{-1} CH_3NH_3Cl and 30.0 mL of 1 mol L^{-1} CH_3NH_2

Calculate the pH of the resultant buffer solution. p*K*a ($CH_3NH_3^+$) = 10.64

(ii) Explain the effect on the solution formed in (i) when a small amount of acid is added.

2014: 1c: An aqueous solution containing a mixture of HF and sodium fluoride, NaF, can act as a buffer solution.

Calculate the mass of NaF that must be added to 150 mL of 0.0500 mol L^{-1} HF to give a buffer solution with a pH of 4.02.

Assume there is no change in volume.

M(NaF) = 42.0 g mol^{-1} pK_a(HF) = 3.17

2017: 2a (ii) : Dilute hydrochloric acid, HCl, is added to the NH_3 solution until the ratio of NH_3 to NH_4^+ in the solution is 5:1.

Determine the pH of this solution, and evaluate its ability to resist a change in pH when small volumes of strong acid or base are added.

$$pH = pK_a - \log \frac{[HA]}{[A^-]} \quad \text{or} \quad pH = pK_a + \log \frac{[A^-]}{[HA]}$$

Writing Excellence answers to **Solubility of sparingly soluble salts** questions

Solubility of sparingly soluble salts QUESTION
Question: Silver carbonate, Ag_2CO_3, is a sparingly soluble salt. $K_s(Ag_2CO_3) = 8.10 \times 10^{-12}$ at 25ºC $M(Ag_2CO_3) = 276$ g mol^{-1} (a) Write the solubility product expression, K_s, for silver carbonate (Ag_2CO_3). (b) Calculate the mass of Ag_2CO_3 that will dissolve in 50 mL of water to make a saturated solution at 25ºC.

ANSWER	
1. write the equation for the dissociation of salt	$Ag_2CO_{3(s)} \leftrightharpoons 2Ag^+_{(aq)} + CO_3^{2-}{}_{(aq)}$
2. Write the solubility product expression, K_s, for the salt	$K_s = [Ag^+]^2[CO_3^{2-}]$
3. calculate the solubility, s 2:1 salt Let s = solubility $K_s = 4s^3$ *3sgf and units*	$K_s = 4s^3$ $s = \sqrt[3]{K_s/4}$ $s = 1.27 \times 10^{-4}$ mol L^{-1}
4. calculate number of moles $n = c \times v$ *3sgf and units*	$n = c \times v$ $n = 1.27 \times 10^{-4}$ mol L^{-1} x 0.0500L $n = 6.33 \times 10^{-6}$ *mol*
5. calculate mass of salt $m = n \times M$ *3sgf and units*	$m = n \times M$ $m = 6.33 \times 10^{-6}$ *mol x* 276 g mol^{-1} $m = 1.75 \times 10^{-3}$ g

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Writing Excellence answers to **Common Ion Effect** questions

Common Ion Effect QUESTION	
Question: Show, by calculation, that a precipitate of lead(II) hydroxide, $Pb(OH)_2$, will form when 25.0 mL of a sodium hydroxide solution, NaOH, at pH 12.6 is added to 25.0 mL of a $[Pb^{2+}] = 0.5 \times 0.00421 = 2.105 \times 10^{-3}$ lead(II) nitrate, $Pb(NO_3)_2$, solution. $K_s(Pb(OH)_2) = 8.00 \times 10^{-17}$ at 25°C	
ANSWER	
1. write the equation for the dissociation of salt	$Pb(OH)_2 \rightleftharpoons Pb^{2+} + 2OH^-$
2. Write the solubility product expression, *Q*, for the salt (K_s)	$Q = [Pb^{2+}][OH^-]^2$
3. calculate the solubility, s for the first ion after dilution $[Pb^{2+}] = \frac{c \times v}{\text{total } v}$ *3sgf and units*	$[Pb^{2+}] = \frac{0.00421 \times 0.0250}{0.0500}$ $[Pb^{2+}] = 2.105 \times 10^{-3}$ molL^{-1}
4. calculate the concentration of [OH-] from pH $[OH^-] = 10^{-(14-pH)}$ *3sgf and units*	$[OH^-] = 10^{-(14-pH)}$ $[OH^-] = 10^{-1.4}$ $[OH^-] = 0.0398$ molL^{-1}
5. calculate the solubility, s for the second ion after dilution $[OH^-] = \frac{c \times v}{\text{total } v}$ *3sgf and units*	$[OH^-] = \frac{0.00398 \times 0.0250}{0.0500}$ $[OH^-] = 1.99 \times 10^{-2}$ molL^{-1}
6. Calculate Q from expression $Q = [ion1] \times [ion2]^2$ *3sgf (has no units)*	$Q = [ion1] \times [ion2]^2$ $Q = (2.105 \times 10^{-3}) \times (1.99 \times 10^{-2})^2$ $Q = 8.34 \times 10^{-7}$
7. compare Q and Ks and state whether a precipitate will form or not	$K_s(Pb(OH)_2) = 8.00 \times 10^{-17}$ at 25°C $Q = 8.34 \times 10^{-7}$ Since $Q > K_s$, a precipitate of $Pb(OH)_2$ will form.

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Writing Excellence answers to **Solubility and Equilibrium** questions

Solubility and Equilibrium QUESTION	
Question: The solubility of zinc hydroxide, $Zn(OH)_2$, can be altered by changes in pH. Some changes in pH may lead to the formation of complex ions, such as the zincate ion, $[Zn(OH)_4]^{2-}$ Use equilibrium principles to explain why the solubility of zinc hydroxide increases when the pH is less than 4 or greater than 10.	
ANSWER	
1. write the **equation** for the dissociation of salt	$Zn(OH)_{2(s)} \rightleftharpoons Zn^{2+}_{(aq)} + 2OH^-_{(aq)}$
2. Explain that OH- ions are formed during dissociation	When $Zn(OH)_{2(s)}$ dissolves then $OH^-_{(aq)}$ ions are produced
3. write the **equation** for the reaction of H_3O^+ ions + OH^- ions when adding acid (due to pH being less than 4)	$H_3O^+ + OH^- \rightarrow H_2O$
4. link removal of OH^- ions (product) to **equilibrium** shifting AND change in solubility	When the pH is less than 4 there are excess H_3O^+ ions present. These react with the OH^- ions to produce water and remove OH^- ions from the solution. so equilibrium shifts to the right to produce more $[OH^-]$, therefore more $Zn(OH)_2$ will dissolve, and increase solubility
5. write the **equation** for the formation of the complex ion $[Zn(OH)_4]^{2-}$ with excess OH^- ions (due to pH being greater than 10)	$Zn(OH)_2(s) + 2OH^- \rightarrow [Zn(OH)_4]^{2-}$ OR $Zn^{2+} + 4OH^- \rightarrow [Zn(OH)_4]^{2-}$
6. link removal of OH^- ions (product) to **equilibrium** shifting AND change in solubility	When the pH is greater than 10 there are excess OH^- ions present. These react with the $Zn(OH)_2$ (Zn^{2+}) to produce a soluble complex ion, $[Zn(OH)_4]^{2-}$ and remove OH^- ions from the solution. so equilibrium shifts to the right to produce more $[OH^-]$, therefore more $Zn(OH)_2$ will dissolve, and increase solubility

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Writing Excellence answers to **Concentration of Species** questions

Concentration of Species QUESTION
Question: Ethyl ammonium chloride, $CH_3CH_2NH_3Cl$, is a weak acid that will also react with water. List all the species present in a solution of $CH_3CH_2NH_3Cl$, in order of decreasing concentration. Do not include water. Justify the order you have given. Include equations, where necessary.

ANSWER	
1. write the **equation** for the dissociation of salt	$CH_3CH_2NH_3Cl \rightarrow CH_3CH_2NH_3^+ + Cl^-$
2. link to complete dissociation AND formation of an **(spectator) ion** that does not react further so will be in greatest concentration	$CH_3CH_2NH_3Cl$ completely dissociates so there will be none remaining. The chloride ion does not react further with water and so will be in the greatest concentration.
3. write the **equation** for the weak acid (formed from equation above) in water	$CH_3CH_2NH_3^+ + H_2O \leftrightharpoons CH_3CH_2NH_2 + H_3O^+$
4. link to partial dissociation due to being a **weak acid** AND most will remain so will be next in concentration	The ethanamine ion ($CH_3CH_2NH_3^+$)will react further with water in an acid-base reaction, but only partially as it is a weak acid, leaving it the next in the series.
5. Explain **H_30^+ ions** are formed during reaction in same quantity as conjugate PLUS small contribution from water AND so will be next in concentration	For every mole of $CH_3CH_2NH_3^+$ that reacts with water, 1 mole of $CH_3CH_2NH_2$ and H_3O^+ are formed. However, H_3O^+ is slightly more concentrated than $CH_3CH_2NH_2$, as there is a small contribution from water $H_2O \leftrightharpoons H_3O^+ + OH^-$
6. Explain **conjugate base** are formed during reaction in same quantity as H_30^+ AND so will be next in concentration (but both H_30^+ ions and conjugate will be at smaller concentration to acid as only weak acid)	Next in concentration is $CH_3CH_2NH_2$. Both $CH_3CH_2NH_3^+$ and H_3O^+ will be at lower concentration than $CH_3CH_2NH_3^+$ due to it being a weak acid
7. Finally Explain **OH- ions** present in small amounts from water dissociation only AND so will be last in concentration	OH^- is present in the lowest concentration as this comes from the dissociation of water only.
8. list **species in order**	$Cl^- > CH_3NH_3^+ > H_3O^+ > CH_3NH_2 > OH^-$

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Writing Excellence answers to **Conductivity and Ions** questions

Conductivity and Ions QUESTION

Question: The table shows the pH and electrical conductivity of three solutions. The concentrations of the solutions are the same. Compare and contrast the pH and electrical conductivity of these three solutions. Include appropriate equations in your answer.

Solution	NaOH	CH_3NH_2	CH_3COONa
pH	13.2	11.9	8.98
Electrical conductivity	good	poor	good

ANSWER

1. Identify each solution as either being a **weak or strong acid or base (or salt)** linked to the **pH** (and presence of ions)	**NaOH** is an ionic solid that is a strong base (pH 13.2) **CH_3NH_2** is a weak base (pH 11.9) **CH_3COONa** is an ionic solid that dissociates completely in H_2O. The CH_3COO^- ion is a weak base (pH 8.98)
2. State requirements for **conductivity**	In order to conduct electricity there needs to be the presence of **free moving charged particles**. The more charged particles there are available the better conductivity there will be. Ions in solution provide the charged particle.
3. **Solution NaOH** (pH 13.2) Write equation **AND** link ions formed to conductivity and level of dissociation	$NaOH \rightarrow Na^+ + OH^-$ **NaOH** is an ionic solid that is a strong base and dissociates completely to produce a high OH^- concentration (low $[H_3O^+]$).
4. pH **Solution NaOH** (pH 13.2) Link amounts of H_3O_+ / OH^- ions to pH	Since $[OH^-]$ is high / $[H_3O^+]$ is low, the pH is high (pH13.2)
5. **Solution CH_3NH_2** (pH 11.9) Write equation **AND** link ions formed to conductivity and level of dissociation	$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$ **CH_3NH_2** is a weak base that partially reacts / dissociates / ionises with H_2O producing a lower concentration of OH^-,
6. pH **Solution CH_3NH_2** (pH 11.9) Link amounts of H_3O_+ / OH^- ions to pH (compared to previous solution)	Since $[OH^-]$ is higher than $[H_3O^+]$ the pH is above 7 (pH11.9) but it still has a lower pH than NaOH:
7. **Solution CH_3COONa** (pH 8.98) Equation 1. [salt dissociation] Write equation **AND** link ions formed to conductivity and level of dissociation	$CH_3COONa \rightleftharpoons CH_3COO^- + Na^+$ The **CH_3COONa** is an ionic solid that dissociates completely in H_2O. with a high ion concentration
8. **Solution CH_3COONa** (pH 8.98) Equation 2.[acid reaction] Write equation **AND** link ions formed to conductivity and level of dissociation	$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$ The CH_3COO^- ion is a weak base that partially reacts / dissociates / ionises with H_2O producing a lower concentration of OH^-.
9. pH **Solution H_3COONa** (pH 8.98) Link amounts of H_3O_+ / OH^- ions to pH (compared to previous solution)	There are slightly more $[OH^-]$ ions than $[H_3O^+]$ The pH is closer to 7,(pH 8.98) showing it is the weakest base. Therefore it has a lowest pH

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Writing Excellence answers to **pH Calculations** questions

pH Calculations QUESTION 1. (4 steps excellence)	
Question: Calculate the pH of a 0.109 mol L^{-1} solution of ethanamine. $pK_a(CH_3CH_2NH_3^+) = 10.6$ $K_w = 1.00\times10^{-14}$	
ANSWER	
1. determine if the solution is acid or base (will it accept or donate H^+) – strong or weak And write down all available information	Ethanamine is a weak base c(ethanamine) = 0.109 mol L^{-1} $pK_a(CH_3CH_2NH_3^+) = 10.6$ $K_w = 1.00\times10^{-14}$
2. convert pK_a to K_a $K_a = 10^{-pKa}$	$K_a = 10^{-pKa}$ $K_a = 10^{-10.6}$ $K_a = 2.51 \times 10^{-11}$
3. calculate $[H_3O^+]$ $[H_3O^+] = \sqrt{\frac{Ka \times Kw}{[base]}}$ *3sgf and units*	$[H_3O^+] = \sqrt{\frac{Ka \times Kw}{[base]}}$ $[H_3O^+] = \sqrt{\frac{2.51 \times 10^{-11} \times 1.00\times10^{-14}}{0.109 \text{ mol } L^{-1}}}$ $[H_3O^+] = 1.52\times10^{-12}$ molL^{-1}
4. calculate pH pH = -log $[H_3O^+]$ *3sgf* *Double check answer against expected pH for your solution*	pH = -log $[H_3O^+]$ pH = -log $[1.52\times10^{-12}$ mol$L^{-1}]$ pH = 11.8 *(pH range for weak base is 8-12) yes*
pH Calculations QUESTION 2. (3 steps Merit)	
Question: Calculate the pH of 0.0152 mol L^{-1} CH_3NH_3Cl solution. $K_a(CH_3NH_3^+) = 2.29 \times 10^{-11}$	
ANSWER	
1. determine if the solution is acid or base (will it accept or donate H^+) – strong or weak And write down all available information	CH_3NH_3Cl is a weak acid (salt) c(CH_3NH_3Cl) = 0.0152 mol L^{-1} $K_a(CH_3NH_3^+) = 2.29 \times 10^{-11}$
2. calculate $[H_3O^+]$ $[H_3O^+] = \sqrt{Ka \times c(HA)}$ *3sgf and units*	$[H_3O^+] = \sqrt{Ka \times c(HA)}$ $[H_3O^+] = \sqrt{2.29 \times 10^{-11} \times 0.0152 \text{ mol } L^{-1}}$ $[H_3O^+] = 5.90 \times 10^{-7}$ mol L^{-1}
3. calculate pH pH = -log $[H_3O^+]$ *3sgf* *Double check answer against expected pH for your solution*	pH = -log $[H_3O^+]$ pH = -log $[5.90 \times 10^{-7}$ mol $L^{-1}]$ pH = 6.23 *(pH range for weak acid is 3-6.9) yes*

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Writing Excellence answers to **Titration Curve – Start pH** questions

Titration Curve – Start pH QUESTION
Question: A titration was carried out by adding hydrobromic acid, HBr, to 20.0 mL of aqueous methylamine, CH_3NH_2, solution. The equation for the reaction is: $CH_3NH_2 + HBr \rightarrow CH_3NH_3^+ + Br^-$ $K_a(CH_3NH_3^+) = 2.29 \times 10^{-11}$ $K_w = 1.00\times10^{-14}$ The aqueous methylamine, CH_3NH_2, solution has a pH of 11.8 before any HBr is added. Show by calculation that the concentration of this solution is 0.0912 mol L^{-1}.

ANSWER	
1. determine if starting solution is acid or base (will it accept or donate H^+) – strong or weak And write down all available information	CH_3NH_2 is a weak base pH = 11.8 $K_a(CH_3NH_3^+) = 2.29 \times 10^{-11}$
2. calculate $[H_3O^+]$ $[H_3O^+] = 10^{-pH}$ *3sgf and units*	$[H_3O^+] = 10^{-pH}$ $[H_3O^+] = 10^{-11.8}$ $[H_3O^+] = 1.58 \times 10^{-12}$ molL^{-1}
3. write out K_a expression $K_a = \frac{[base][H_3O^+]}{[conj\ acid]}$ And then $K_a = \frac{[base][H_3O^+]}{[OH^-]}$	$K_a = \frac{[base][H_3O^+]}{[conj\ acid]}$ $K_a = \frac{[CH_3NH_2][H_3O^+]}{[CH_3NH_3^+]}$ And $K_a = \frac{[CH_3NH_2][H_3O^+]}{[OH^-]}$
4. rearrange to calculate $[CH_3NH_2]$ $[CH_3NH_2] = \frac{K_a \times K_w}{[H_3O^+]^2}$ *Assumptions: [conj] = $[H_3O^+]$* *$[OH^-] = K_w / [H_3O^+]$* *3sgf and units*	$[CH_3NH_2] = \frac{K_a \times K_w}{[H_3O^+]^2}$ *Assumptions: [base] = $[H_3O^+]$* *$[OH^-] = K_w / [H_3O^+]$* $[CH_3NH_2] = \frac{2.29 \times 10^{-11} \times 1.00\times10^{-14}}{(1.58 \times 10^{-12}\ mol L^{-1})^2}$ $[CH_3NH_2] = 0.0912$ mol L^{-1}

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Writing Excellence answers to **Titration Curve – After the Start pH** questions

Titration Curve – after the Start pH QUESTION
Question: : 20.0 mL of 0.0896 mol L^{-1} ethanoic acid is titrated with 0.100 mol L^{-1} sodium hydroxide. p*K*a (CH_3COOH) = 4.76 Calculate the pH of the titration mixture after 5.00 mL of NaOH has been added. $K_w = 1 \times 10^{-14}$

ANSWER	
1. determine if starting solution is acid or base (will it accept or donate H^+) – strong or weak And write down all available information	Starting solution of ethanoic acid is weak acid v(CH_3COOH) = 20mL = 0.0200L c(CH_3COOH) = 0.0896 mol L^{-1} v(NaOH) = 5mL = 0.00500L c(NaOH) = 0.100 mol L^{-1} total volume = 25mL = 0.0250L
2. Write down neutralisation equation	$NaOH(aq) + CH_3COOH_{(aq)} \rightarrow NaCH_3COO_{(aq)} + H_2O_{(l)}$
3. calculate *n*(CH_3COOH at start) n = cv *3sgf and units*	n = cv n = 0.0896 × (20 ×10^{-3}) n = 1.79 × 10^{-3} mol
4. calculate *n*(NaOH) and therefore n(CH_3COO-) n = cv *assume n*(NaOH) = *n*(CH_3COO-) *3sgf and units*	n = cv n = 0.1 × (5 × 10^{-3}) n = 5.00 × 10^{-4} mol *assume n*(NaOH) = *n*(CH_3COO-) *n*(CH_3COO-) = 5.00 × 10^{-4}
5. calculate *n*(CH_3COOH) After 5 mL NaOH added: (total 25mL) =n(CH_3COOH – n(CH_3COO-) after 5mL *3sgf and units*	*n*(CH_3COOH) After 5 mL = n(CH_3COOH – n(CH_3COO-) after 5mL *n*(CH_3COOH) After 5 mL = (1.79 × 10^{-3} mol – 5.00 × 10^{-4} mol) *n*(CH_3COOH) After 5 mL = 1.29 × 10^{-3} mol
6. calculate c(CH_3COOH) c = n/v *3sgf and units*	c = n/v c = 1.29 × 10^{-3} mol / 0.0250L c = 0.0516 mol L^{-1}
7. calculate c(CH_3COO^-) c = n/v (remember v = 25mL) *3sgf and units*	c = n/v c = 5.00 × 10^{-4} mol / 0.0250L c = 0.0200 mol L^{-1}
8. Calculate pH $pK_a = 4.76$ $pH = pK_a + \log \frac{[CH_3COO^-]}{[CH_3COOH]}$ *3sgf* *Check pH against estimate on curve*	$pH = pK_a + \log \frac{[CH_3COO^-]}{[CH_3COOH]}$ $pH = 4.75 + \log \frac{[0.0200]}{[0.0515]}$ pH = 4.35

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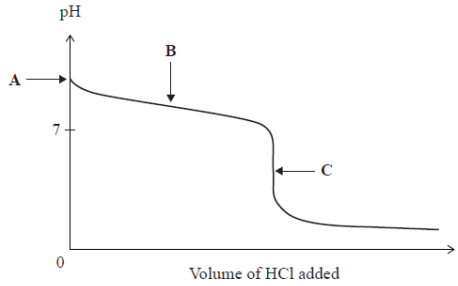
Writing Excellence answers to **Titration Curve – Equivalence Point pH** questions

Titration Curve – Equivalence Point pH QUESTION
Question: 20.00 mL of 0.320 mol L^{-1} ammonia, NH_3, is titrated with 0.640 mol L^{-1} hydrochloric acid, HCl. p$K_a(NH_4^+)$ = 9.24 Show, by calculation, that the pH at the equivalence point (point C) is 4.96. $K_w = 1 \times 10^{-14}$ 

ANSWER	
1. determine if equivalence point is greater or less than 7 (from curve or strong base/weak acid strong acid/weak base And write down all available information	ammonia, NH_3 is weak base and hydrochloric acid, HCl is strong acid so equivalence point <7 v(NH_3) = 20.00 mL = 0.0200L c(NH_3) = 0.320 mol L^{-1} c(HCl) = 0.640 mol L^{-1} p$K_a(NH_4^+)$ = 9.24
2. Write down neutralisation equation	$NH_3 + HCl \rightarrow NH_4^+ + Cl^-$
3. calculate *n*(Base) to neutralise (and reach equivalence point and therefore n(Acid) from 1:1 equation) n = cv *also assume* *n*(NH_3) = *n*(NH_4^+) *3sgf and units*	n(NH_3) = cv n(NH_3) = 0.320 mol L^{-1} x 0.0200L n(NH_3) = 6.40×10^{-3} mol
4. calculate *v*(Acid) to neutralise (n(NH_3) = n(HCl) from 1:1 equation) v=n/c *3sgf and units*	v=n/c v=6.40×10^{-3} mol / 0.640 mol L^{-1} v=0.0100L (10.0mL)
5. calculate [B^+] c = n/total v *also assume* *n*(B) = *n*(B^+) see step 3. B = NH_3 B+ = NH_4^+ total v = start volume base + v acid added *3sgf and units*	c = n/total v c = 6.40×10^{-3} mol /0.0300L c = 0.213 molL^{-1}
6. calculate [H_3O^+] $K_a = 10^{-pKa}$ [H_3O^+] = $\sqrt{Ka \times c(B^+)}$ *3sgf and units* $B^+ = HA$	[H_3O^+] = $\sqrt{Ka \times c(B^+)}$ [H_3O^+] = $\sqrt{10^{-9.24} \times 0.213}$ molL^{-1} [H_3O^+] = 1.11×10^{-5} molL^{-1}
7. Calculate pH pH = -log [H_3O^+] *3sgf* *Check pH against estimate on curve*	pH = -log [H_3O^+] pH = -log [1.11×10^{-5} molL^{-1}] pH = 4.96

NOTE: The white column is how your answer would appear on your test paper so make sure you **write out complete sentences**. The grey area is just to help you structure your answer and would not appear in the question.

Writing Excellence answers to **Buffer pH Calculation** questions

Buffer pH Calculation QUESTION	
Question: The following two solutions from part (a) are mixed to form a buffer solution: 20.0 mL of 1 mol L^{-1} CH_3NH_3Cl and 30.0 mL of 1 mol L^{-1} CH_3NH_2 Calculate the pH of the resultant buffer solution. p*K*a ($CH_3NH_3^+$) = 10.64 $K_w = 1 \times 10^{-14}$	
ANSWER	
1. Write out Ka expression	$K_a = \frac{[CH_3NH_2][H_3O^+]}{[CH_3NH_3^+]}$
2. rearrange expression to calculate $[H_3O^+]$	$[H_3O^+] = \frac{K_a [CH_3NH_3^+]}{[CH_3NH_2]}$ Or pH = pKa + log $\frac{[CH_3NH_3^+]}{[CH_3NH_2]}$
3. calculate $[CH_3NH_2]$ $[CH_3NH_2] = \frac{v(CH_3NH_2) \times c(CH_3NH_2)}{\text{total volume}}$ *3sgf and units*	$[CH_3NH_2] = \frac{v(CH_3NH_2) \times c(CH_3NH_2)}{\text{total volume}}$ $[CH_3NH_2] = \frac{0.0300L \times 1.00molL^{-1}}{0.0500L}$ $[CH_3NH_2] = 0.600molL^{-1}$
4. calculate $[CH_3NH_3^+]$ $[CH_3NH_3^+] = \frac{v(CH_3NH_3^+) \times c(CH_3NH_3^+)}{\text{total volume}}$ *3sgf and units*	$[CH_3NH_3^+] = \frac{v(CH_3NH_3^+) \times c(CH_3NH_3^+)}{\text{total volume}}$ $[CH_3NH_3^+] = \frac{0.0200L \times 1.00molL^{-1}}{0.0500L}$ $[CH_3NH_3^+] = 0.400molL^{-1}$
5. calculate *pH* $pH = pKa + \log \frac{[A-]}{[HA]}$ *3sgf*	$pH = pKa + \log \frac{[A-]}{[HA]}$ $pH = pKa + \log \frac{0.400molL^{-1}}{0.600molL^{-1}}$ *pH = 10.8*

NOTE: The white column is how your answer would appear on your test paper so make sure you **write out complete sentences**. The grey area is just to help you structure your answer and would not appear in the question.