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)

All NCEA answers can be found on C3.6 ppt

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.

Summary of solubility calculations – Q1

Question: Calculate K_s	Question: Calculate mass
Starting information: Mass, Molar mass and Volume	**Starting information:** Ks and Molar Mass
Calculate c	Calculate s
n = m/M and then c = n/v	1:1 $s = \sqrt{Ks}$ 2:1 $s = \sqrt[3]{Ks/4}$

Assume concentration (c) = s

Calculate Ks	Calculate Mass
1:1 salt Ks = [s][s] (or s^2)	use n = c.v
2:1 salt Ks = $[2s]^2[s]$ (or $4s^3$)	then m = n x M

Mass (g) concentration ($molL^{-1}$) answers to 3 sgf

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 25°C $M(Ag_2CO_3) = 276$ 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 25°C
$M(Ag_2CO_3) = 276$ g mol^{-1} Calculate the mass of Ag_2CO_3 that will dissolve in 50 mL of water to make a saturated solution at 25°C.

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.
$K_s(Cu(OH)_2) = 4.80 \times 10^{-20}$

2018: 1b (i). Write the equation for the equilibrium occurring in a saturated solution of calcium fluoride, Ca F_2.
(ii). Calculate the solubility of Ca F_2 in water at 25°C.
$K_s(Ca\ F_2) = 3.20 \times 10^{-11}$

2019: Question: 1a (i). Write the equation for the equilibrium occurring in a saturated solution of zinc hydroxide, $Zn(OH)_2$.

(ii) Write the expression for $K_s(Zn(OH)_2)$.

(iii) Calculate the solubility of $Zn(OH)_2$ in water at 25°C, and give the $[Zn^{2+}]$ and $[OH^-]$ in the solution.

$K_s(Zn(OH)_2) = 3.80 \times 10^{-17}$

2020: Question: 2a (i). Write the equation for the equilibrium occurring in a saturated solution of lead bromide, $PbBr_2$.

(ii) Write the expression for $K_s(PbBr_2)$.

(iii) Calculate the solubility of $PbBr_2$ in water at 25 °C. $K_s(PbBr_2) = 2.10 \times 10^{-6}$

2020: Question: 2c (ii). Calculate the solubility of a saturated solution of nickel hydroxide at pH 8.25.

$K_s(Ni(OH)_2) = 6.00 \times 10^{-16}$

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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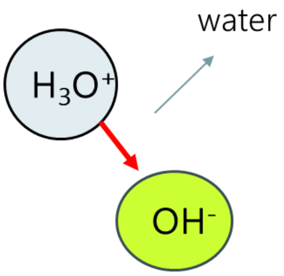
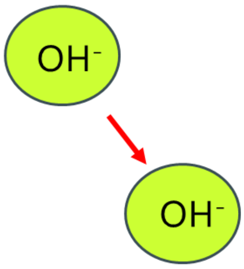
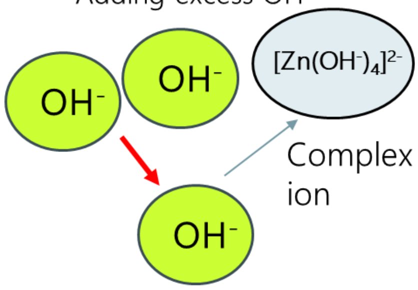
Summary of Equilibrium changes in Solubility – Q2

$Zn(OH)_{2(s)} \rightleftharpoons Zn^{2+}{}_{(aq)} + 2OH^{-}{}_{(aq)}$)

pH is below 4	pH is above 4 but below 10	pH is above 10
	Adding few drops OH^- 	Adding excess OH^- 
Hydronium ions in acid solution **neutralise** OH-, and removing as product.	Total OH- concentration increases, therefore more product.	Complex ion forms, and removes product., by removing Zn^{2+} ions.
Forward reaction increases	Reverse reaction increases	Forward reaction increases
Solubility increases	Solubility decreases	Solubility increases

Remove product – solubility increases. Add to product – solubility decreases

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) 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.

2018: 1c. Explain the effect of the following on the solubility of iron(III) hydroxide, $Fe(OH)_3$, in water.
Include relevant equation(s) in your answer. No calculations are necessary.
(i) pH lowered below 4

2018: 1c. Explain the effect of the following on the solubility of iron(III) hydroxide, $Fe(OH)_3$, in water.
Include relevant equation(s) in your answer. No calculations are necessary.
(ii) Potassium thiocyanate, KSCN, solution added

2018: (a) (i) Write the solubility product expression, *K*s, for silver chloride, AgCl.

(ii) Why does the solubility of AgCl decrease when a small volume of silver nitrate, $AgNO_3$, solution is added to a saturated solution of AgCl?
Explain your answer.

2019: Question: 1b Use equilibrium principles to explain why the solubility of $Zn(OH)_2$ increases when an excess of dilute sodium hydroxide, NaOH, is added.
Include relevant equation(s) in your answer.

2020: Question: 2c (i) Explain the effect of the following on the solubility of nickel hydroxide, $Ni(OH)_2$, in water.

Include relevant equation(s) in your answer. No calculations are necessary.
• Ammonia solution, $NH_{3(aq)}$, is added:
• The pH is decreased below 4:

https://www.sciencesource.com/archive/Silver-chloride-precipitate-SS21845965.ht

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] x $[ion2]^2$ *3sgf (has no units)*	
7. compare Q and Ks and state whether a precipitate will form or not	

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Summary of Common ion calculations – Q3

Q = ratio of the concentrations of products and reactants.

1:1 salt	2:1 salt
Q = (c) x (c)	Q = (c) x (c)^2

Where can c come from?

s (calculated from K_s)

c provided in the question

n = m/M and then c = n/v

C(OH^-) from pH

pOH = 14 - pH [OH-] = 10^{-pOH}

c(new) = $\frac{\text{initial volume}}{\text{total volume}}$ x conc.

If I.P. < K_s NO precipitate If I.P. > K_s precipitate forms

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.

2018: 3a (iii). Show by calculation whether a precipitate of AgCl will form when 70.0 mL of 0.0220 mol L^{-1} $AgNO_3$ is added to 50.0 mL of 0.0550 mol L^{-1} sodium chloride, NaCl.

$K_s(AgCl) = 1.80 \times 10^{-10}$

2019: Question: 1a (iv). The presence of a common ion decreases the solubility of a sparingly soluble solid, such as $Zn(OH)_2$.

Calculate the concentration of the hydroxide ions, OH^-, in solution after 25.0 mL of 0.210 mol L^{-1} zinc chloride, $ZnCl_2$, solution was added to 25.0 mL of a saturated $Zn(OH)_2$ solution.

$K_s(Zn(OH)_2) = 3.80 \times 10^{-17}$

2019: Question: 1c Determine whether a precipitate of $Zn(OH)_2$ will form when 30.0 mL of sodium hydroxide solution, NaOH, at pH 13.1 is added to 20.0 mL of 0.0242 mol L^{-1} zinc nitrate, $Zn(NO_3)^2$.

$K_s(Zn(OH)_2) = 3.80 \times 10^{-17}$

2020: Question: 2b Determine whether a precipitate of lead bromide, $PbBr_2$, will form when 125 mL of 0.0365 mol L^{-1} lead nitrate, $Pb(NO_3)_2$, is added to 175 mL of 0.00262 mol L^{-1} magnesium bromide, $MgBr_2$.

. $K_s(PbBr_2) = 2.10 \times 10^{-6}$

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.

Summary of Dissociation equations and Equilibrium expressions - Q4

Weak Acids	Weak Bases
HF, CH_3COOH, and NH_4^+	NH_3, CH_3NH_2, and CH_3COO^-
Write an equation using:	Write an equation using:
$HA_{(aq)} + H_2O_{(l)} \leftrightharpoons H_3O^+_{(aq)} + A^-_{(aq)}$	$B_{(aq)} + H_2O_{(l)} \leftrightharpoons OH^-_{(aq)} + BH^+_{(aq)}$

Assume H_2O has no change in concentration

Write an expression using:	Write an expression using:
$K_a = \dfrac{[H_3O^+][A^-]}{[HA]}$	$K_b = \dfrac{[OH^-][BH^+]}{[B]}$

strong acids HCl, HBr, HNO_3, H_2SO_4 strong bases KOH, NaOH

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.

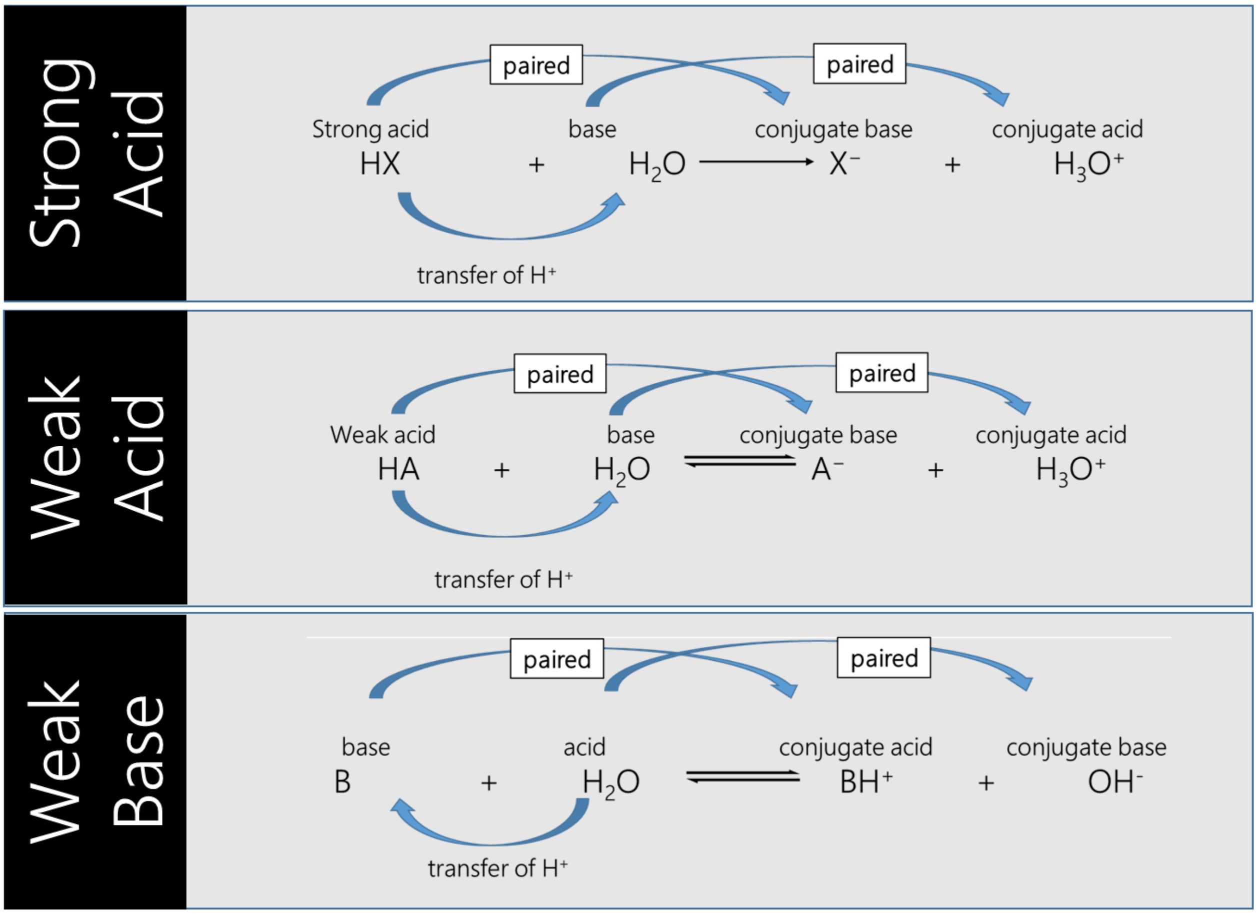
2016: 2a: (i) Ethanamine, $CH_3CH_2NH_2$, is a weak base.
pK_a($CH_3CH_2NH_3^+$) = 10.6 K_a($CH_3CH_2NH_3^+$) = 2.51×10^{-11}

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:

Acid and Base Dissociation



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 = 1.00×10^{-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^+] $[H_3O^+] = \sqrt{\dfrac{Ka \times Kw}{[base]}}$ *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($CH_3NH_3^+$) = 2.29×10^{-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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pH Calculations – Q5

Convert pK_a to K_a

$K_a = 10^{-pKa}$

If given Weak Acid	If given Weak Base
HF, CH_3COOH, and NH_4^+	NH_3, CH_3NH_2, and CH_3COO^-
$[H_3O^+] = \sqrt{K_a \times c(HA)}$	$[H_3O^+] = \sqrt{(K_a \times K_w \div [B])}$

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

strong acids HCl, HBr, HNO_3, H_2SO_4 strong bases KOH, $NaOH$

Comparing weak acids - pH, pKa and ions – Q6

Strong acid Weak acid

pKa value 0 13

stronger acid weaker acid

Link to level of acid dissociation

stronger acid	weaker acid
More dissociation of acid to produce more H_3O^+ ions and conjugate base	less dissociation of acid to produce more H_3O^+ ions and conjugate base

Link to pKa and pH and acid strength

stronger acid	weaker acid
pKa is a smaller value, as is the pH, therefore this is a stronger acid	pKa is a larger value, as is the pH, therefore this is a weaker acid

Weak acids CH_3COOH, HF

Past NCEA questions pH Calculations

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.

2018: 1a: (a) When sodium ethanoate, CH_3COONa, is dissolved in water, the resulting solution has a pH greater than 7 due to the production of hydroxide ions, OH^-, as shown in the equation below.

$CH_3COO^- + H_2O \leftrightharpoons CH_3COOH + OH^-$
$pK_a(CH_3COOH) = 4.76$ $K_a(CH_3COOH) = 1.74 \times 10^{-5}$
Calculate the pH of a 0.420 mol L^{-1} CH_3COONa solution.

2019: Question: 3a Two solutions of equal concentration were prepared: one of ethanoic acid, CH_3COOH, and one of ammonium chloride, NH_4Cl.

$pK_a(CH_3COOH) = 4.76$ $pK_a(NH_4^+) = 9.24$
(i) Explain which solution would have the lower pH.
Your answer should refer to the concentration of relevant ion(s) in each solution.
No calculations are necessary

2019: Question: 3a Two solutions of equal concentration were prepared: one of ethanoic acid, CH_3COOH, and one of ammonium chloride, NH_4Cl.

$pK_a(CH_3COOH) = 4.76$ $pK_a(NH_4^+) = 9.24$
(iii) The ethanoic acid solution has a $[H_3O^+]$ of 1.78×10^{-3} mol L^{-1}.
Calculate the concentration of the ethanoic acid solution.

2020: Question: 1c (ii) If the NH_4Cl solution has a pH of 4.70, calculate its concentration.

$K_a(NH_4^+) = 5.75 \times 10^{-10}$ $pK_a(NH_4^+) = 9.24$

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**	

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 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

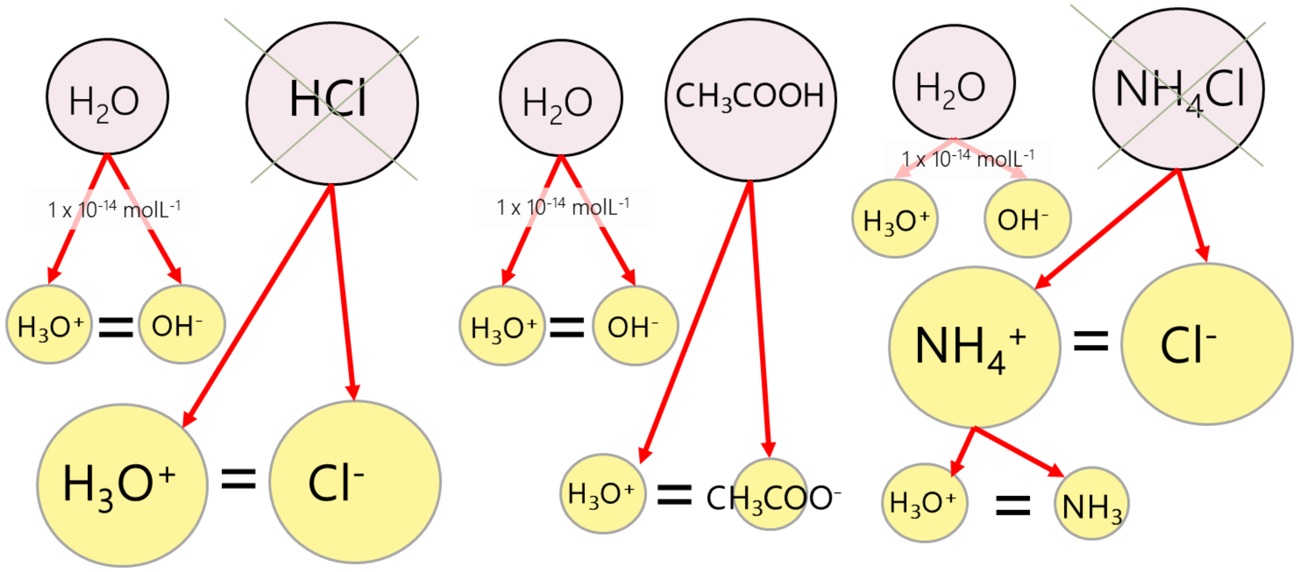
ANSWER

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) <u>Write equation</u> **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) <u>Write equation</u> **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] <u>Write equation</u> **AND** link ions formed to conductivity and level of dissociation	
8. **Solution CH_3COONa** (pH 8.98) Equation 2.[acid reaction] <u>Write equation</u> **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)	

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.

Summary of Species/conductivity in Solution – Acid –Q7

Strong Acids	Weak Acids	Acid salts



Strong Acids: Complete dissociation
No **strong acid** remains

High conductivity

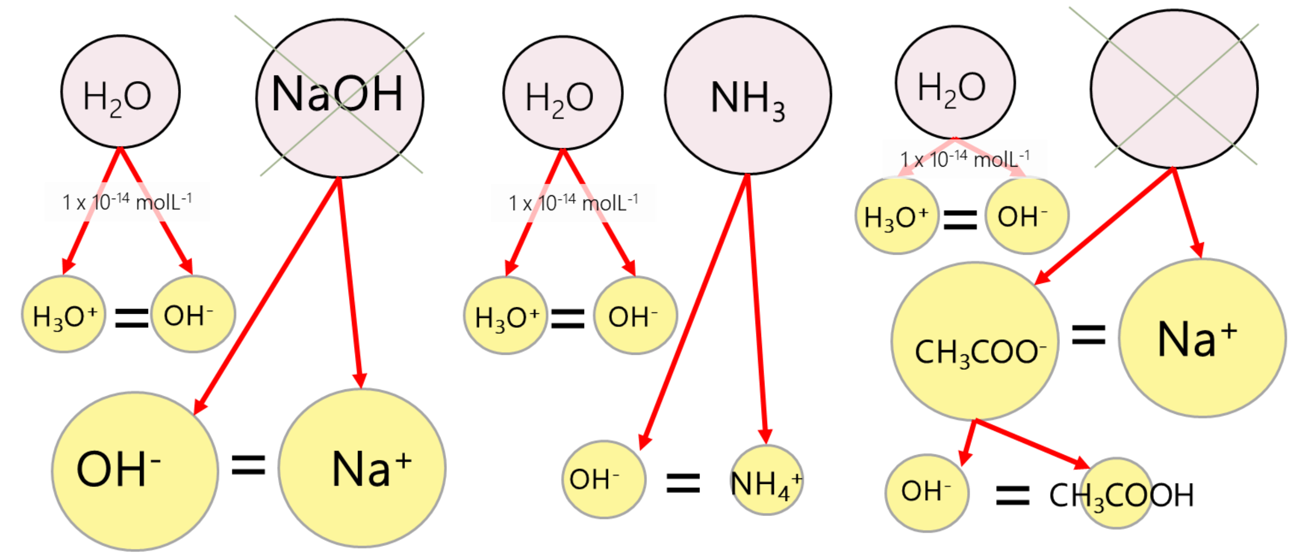
Weak Acids: Partial dissociation
Most weak acid remains

Low conductivity

Water concentration is assumed to remain constant so is left out

Summary of Species / conductivity in Solution – Base – Q7

Strong Bases	Weak Bases	Base salts



Strong Bases: Complete dissociation
No **strong base** remains

High conductivity

Weak Bases: Partial dissociation
Most weak base remains

Low conductivity

water concentration is assumed to remain constant so is left out

Past NCEA questions Species in Solution and Conductivity

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

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.

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.

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

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: 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.

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

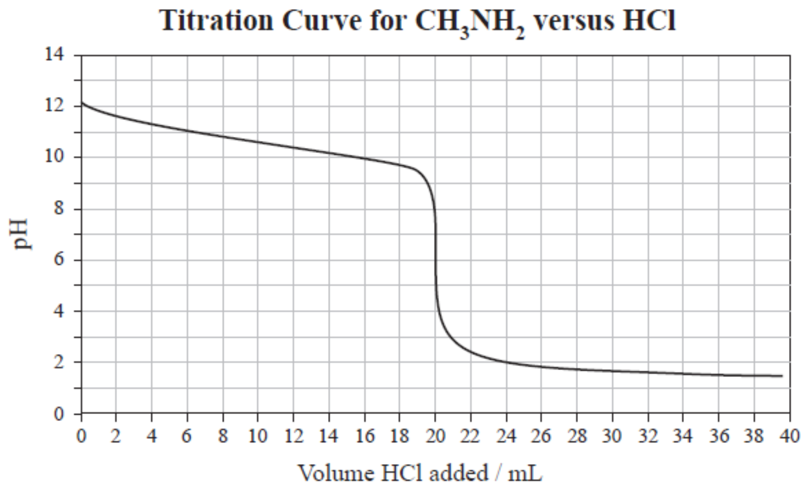
(i) Identify the chemical species present at the equivalence point, other than water.

2018: 2. A titration was carried out by adding 0.210 mol L^{-1} hydrochloric acid, HCl, to 25.0 mL of 0.168 mol L^{-1} methanamine, CH_3NH_2.

The equation for the reaction is: $HCl + CH_3NH_2 \rightarrow CH_3NH_3^+ + Cl^-$

$pK_a(CH_3NH_3^+) = 10.6$ $K_a(CH_3NH_3^+) = 2.51 \times 10^{-11}$

(b) (i) List all the species present in the solution at the equivalence point in order of decreasing concentration. Do not include water.



(c) Why is the solution at the equivalence point a better electrical conductor than the initial solution of methanamine? Your answer should include relevant equation(s) and elaborate on the relative concentrations of the different species in each solution. No calculations are necessary.

2019: Question: 3a (ii) Evaluate the electrical conductivity of the CH_3COOH and NH_4Cl solutions.

Include relevant equation(s) in your answer.

2020: Question: 3b Why are hydrochloric acid and sodium ethanoate solutions both good electrical conductors? Justify your answer, including any relevant equation(s).

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.

Effectiveness of Buffer solutions – Q8

Weak Acids and conjugates

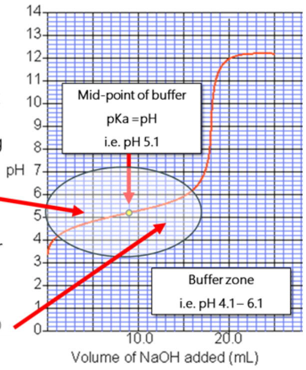
The pH of a solution falls within a buffer zone when it is 1 pH either side of the pKa), and therefore it will function as a buffer.

When the pH < p*K*a, this means [weak acid] > [conjugate], so the buffer will be more effective against added strong base.

$COOH_{(aq)} + OH^-_{(aq)} \leftrightharpoons COO^-_{(aq)} + H_2O_{(l)}$

When the pH > p*K*a, this means [conjugate] > [weak acid], so the buffer will be more effective against added strong acid.

$COO^-_{(aq)} + H_3O^+_{(aq)} \leftrightharpoons COOH_{(aq)} + H_2O_{(l)}$



Weak Bases and conjugates

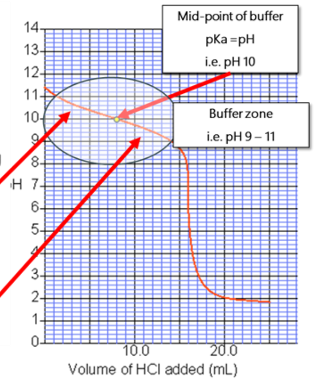
The pH of a solution falls within a buffer zone when it is 1 pH either side of the pKa), and therefore it will function as a buffer.

When the pH > p*K*a, this means [weak base] > [conjugate], so the buffer will be more effective against added strong acid.

$CH_3NH_2 + H_3O^+ \rightarrow CH_3NH_3^+ + H_2O$

When the pH < p*K*a, this means [conjugate] > [weak base], so the buffer will be more effective against added strong base.

$CH_3NH_3^+ + OH^- \rightarrow CH_3NH_2 + H_2O$



Weak acid	Conjugate base
CH_3COOH	CH_3COO^-
$CH_3COOH + H_2O \leftrightharpoons CH_3COO^- + H_3O^+$	
$CH_3COO^- + H_2O \leftrightharpoons CH_3COOH + OH^-$	
HF	F^-
$HF + H_2O \leftrightharpoons F^- + H_3O^+$	
$F- + H_2O \leftrightharpoons HF^- + OH^-$	
NH_4^+	NH_3
$NH_4^+ + H_2O \leftrightharpoons NH_3 + H_3O^+$	
$NH_3 + H_2O \leftrightharpoons NH_4^+ + OH^-$	
$CH_3NH_3^+$	CH_3NH_2
$CH_3NH_3^+ + H_2O \leftrightharpoons CH_3NH_2 + H_3O^+$	
$CH_3NH_2 + H_2O \leftrightharpoons CH_3NH_3^+ + OH^-$	

Pathways for pH calculations

$K_a = 10^{-pKa}$ $K_w = 1 \times 10^{-14}$

$B + H_2O \leftrightharpoons BH^+ + OH^-$

$HA + H_2O \leftrightharpoons A^- + H_3O^+$

pH of Base

$K_b = \frac{[BH^+][OH^-]}{[B]}$

$K_b = K_w / K_a$

$[OH^-] = \sqrt{K_b \times c(B)}$

$[H_3O^+] = K_w/[OH^-]$

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

pH of Acid

$A^- = H_3O^+$ [HA] = c(HA)

$K_a = \frac{[A^-][H_3O^+]}{[HA]}$

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

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

pH of buffer

$A^- = H_3O^+$

$[A-] = \frac{K_a \times [HA]}{10^{-pH}}$

$n = c \times v$

$m = n \times M$

mass

$pH = PK_a + \log \frac{[A-]}{[HA]}$

pH

$[H_3O^+] = 10^{-pH}$

$[H_3O^+]$

Past NCEA questions Buffers

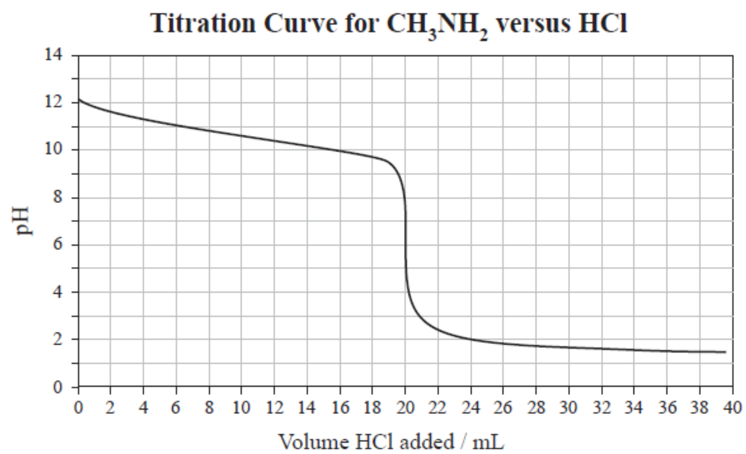
2013:1c: (i) The following two solutions from part (a) are mixed to form a buffer solution:
20.0 mL of 1 mol L^{-1} CH_3NH_3Cl and 30.0 mL of 1 mol L^{-1} CH_3NH_2
Calculate the pH of the resultant buffer solution. pKa ($CH_3NH_3^+$) = 10.64

(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.

2018: 2. A titration was carried out by adding 0.210 mol L^{-1} hydrochloric acid, HCl, to 25.0 mL of 0.168 mol L^{-1} methanamine, CH_3NH_2. The equation for the reaction is: $HCl + CH_3NH_2 \rightarrow CH_3NH_3^+ + Cl^-$
pKa($CH_3NH_3^+$) = 10.6 K_a($CH_3NH_3^+$) = 2.51×10^{-11}



(a) Between pH 9.60 – 11.6, the solution is a buffer.
(i) From the titration curve, estimate the volume of the HCl solution that must be added to the CH_3NH_2 solution above to make a buffer solution of pH 10.0.

(ii) Explain how the buffer solution resists large changes in pH as the HCl solution is added between a pH of 9.60 – 11.6.
Include an appropriate equation in your answer.

2018: 3b. 5.11 g of sodium methanoate, HCOONa, was added to 125 mL of 0.105 mol L^{-1} methanoic acid, HCOOH, to make a buffer solution. Assume there is no change in the total volume.
$pK_a(HCOOH) = 3.74$ $K_a(HCOOH) = 1.82 \times 10^{-4}$
(i) Give the pH range over which the resulting solution will function as a buffer.

(ii) Show, by calculation, that the pH of this buffer solution is 4.50.
$M(HCOONa) = 68.0 \text{ g mol}^{-1}$

2019: Question: 3b (i) Dilute hydrochloric acid, HCl, is added to a solution of sodium ethanoate, CH_3COONa, until the ratio of CH_3COONa to ethanoic acid, CH_3COOH, in the solution is two to five (2:5).

Calculate the pH of this buffer solution. p*K*a(CH3COOH) = 4.76

(ii) Explain why this buffer solution would be more effective at resisting a change in pH when a small volume of strong base is added, rather than strong acid. Dilute hydrochloric acid, HCl, is added to a solution of sodium ethanoate, CH_3COONa
Your answer should include an equation to show how the buffer neutralises added strong base.

(iii) How would the pH of this buffer solution be affected when it is diluted with water? Explain your answer.

2020: Question: 1a An aqueous solution containing a mixture of ammonium chloride, NH_4Cl, and ammonia, NH_3, can act as a buffer solution.

$K_a(NH_4^+) = 5.75 \times 10^{-10}$ $pK_a(NH_4^+) = 9.24$
(i) Give the pH range over which the solution will function as a buffer.

(ii) Explain why the addition of a small volume of nitric acid, HNO_3, to this buffer solution will not result in a significant change in pH.
Your answer should include relevant equation(s).

2020: Question: 1b (i) Calculate the mass of NH_4Cl that must be added to 200 mL of 0.0500 mol L^{-1} NH_3 to give a buffer solution with a pH of 8.75.
Assume there is no change in volume when the solid is added.
$M(NH_4Cl) = 53.5 \text{ g mol}^{-1}$

(ii) Explain whether the buffer in part (i) will be more effective at neutralising strong acid or strong base.

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	
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*	

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 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*	

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.

Summary of steps – After amount base/acid added – Q9

Concentration and volume given for both acid and base

Weak base vs Strong Acid	Weak acid vs Strong Base
1. Calculate n weak base n(weak base (start)) = cv	1. Calculate n weak acid n(weak acid (start)) = cv
2. Calculate n strong acid n(strong acid (start)) = cv	2. Calculate n strong base n(strong base (start)) = cv
After acid added	*After base added*
3. n(weak base) = n(weak base(start)) – n(acid)	3. n(weak acid) = n(weak acid (start)) – n(base)
4. c(weak base) = n(weak base) / total volume	4. c(weak acid) = n(weak acid) / total volume
NOTE: n(conj. acid) = n(strong acid)	NOTE: n(conj. base) = n(strong base)
3. c(conj. acid) = n(conj. acid) / total volume	3. c(conj. base) = n(conj. base) / total volume
$pH = pK_a + \log\left(\frac{[\text{weak base}]}{[\text{conj. acid}]}\right)$	$pH = pK_a + \log\left(\frac{[\text{conj. base}]}{[\text{weak acid}]}\right)$

Calculating pH after an amount of base (or acid) is added

Sample question: 20.0 mL of 0.0896 mol L^{-1} ethanoic acid is titrated with 0.100 mol L^{-1} sodium hydroxide. *pKa* (CH_3COOH) = 4.76

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

$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$

This is the equation that the K_a is derived from – but as most of the CH_3OO^- is produced from the NaOH reacting, this is the concentration we use to calculate pH

$NaOH(aq) + CH_3COOH_{(aq)} \rightleftharpoons NaCH_3COO_{(aq)} + H_2O_{(l)}$

$n(CH_3COOH \text{ at start}) = 0.0896 \times (20 \times 10^{-3}) = 1.79 \times 10^{-3}$ mol

Calculate the number of moles of 20ml acid
n = c x v

$n(NaOH \text{ added}) = 0.1 \times (5 \times 10^{-3}) = 5 \times 10^{-4}$ mol

Calculate the number of moles of 5ml base

After 5 mL NaOH added: (total 25mL)

$n(CH_3COOH) = 1.29 \times 10^{-3}$ mol

Subtract moles of base from acid - neutralised
($n(CH_3COOH – n(NaOH)$ after 5mL)

Remember 1 mol CH_3COO^- produced for each mol NaOH reacting.
Also produced from CH_3COOH reacting with water but K_a is small so can disregard

$n(CH_3COO^-) = 5 \times 10^{-4}$ mol

c=n/v
Volume is 25ml total

$[CH_3COOH] = 0.0516$ mol L^{-1} $c = 1.29 \times 10^{-3} / 0.025$

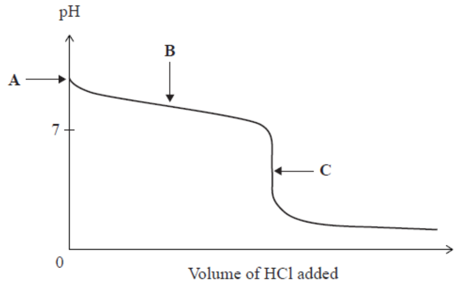
$[CH_3COO^-] = 0.0200$ mol L^{-1} $c = 5 \times 10^{-4} / 0.025$

$pH = pK_a + \log\left(\frac{[CH_3COO^-]}{[CH_3COOH]}\right)$

pH = 4.35

As volume added is less than midpoint of buffer (1/2 equivalence point) will be less than pK_a

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. $pK_a(NH_4^+) = 9.24$ 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*	

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.

Summary of steps - Equivalence point pH – Q10

From neutralisation equation

n(acid) =n(weak base) = n(conjugate acid)

From neutralisation equation

n(base) =n(weak acid) = n(conjugate base)

Weak base vs Strong Acid

1. Calculate n weak base
   n(weak base) = cv (info from question)
2. Calculate v strong acid
   v(strong acid) =n/c
3. Calculate c conjugate acid
   c (conjugate acid)= n/total v
4. $[H_3O^+] = \sqrt{K_a \times c(\text{conjugate acid})}$
5. pH = -log $[H_3O^+]$

Weak acid vs Strong Base

1. Calculate n weak acid
   n(weak acid) = cv (info from question)
2. Calculate v strong base
   v(strong base) =n/c
3. Calculate c conjugate base
   c (conjugate base)= n/total v
4. $[H_3O^+] = \sqrt{\dfrac{K_a \times K_w}{c(\text{conjugate base})}}$
5. pH = -log $[H_3O^+]$

Alternative

$$K_a = \frac{[H_3O^+]^2}{(c(\text{acid}) \times \text{original v/total v})}$$

If given v in the question (from graph)

Summary of steps - Equivalence point pH – Q10

Finish here

pH= -log$[H_3O^+]$ ← Need to find pH

↑

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

↑

Need to find H_3O^+

↑

c (conj acid)= n/total v

You can work also backwards to find out what calculations you need to do

n(base) = n(conj acid) → c (conj acid)= n/total v

Start v(base) + v(Acid) → c (conj acid)= n/total v

Need to find n(conj acid) → n(base) = n(conj acid)

Need to find v(conj acid) → Start v(base) + v(Acid)

Start here

n(base) = c.v → Need to find n(conj acid)

n(base) = n(acid)

v(acid) = n/c → Need to find v(conj acid)

Given conc of acid + base and volume start base

Past NCEA questions Titration Calculations (Part ONE)

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

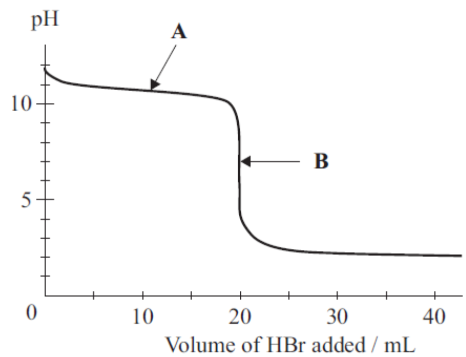
Calculate the pH of the ethanoic acid before any NaOH is added.

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}.

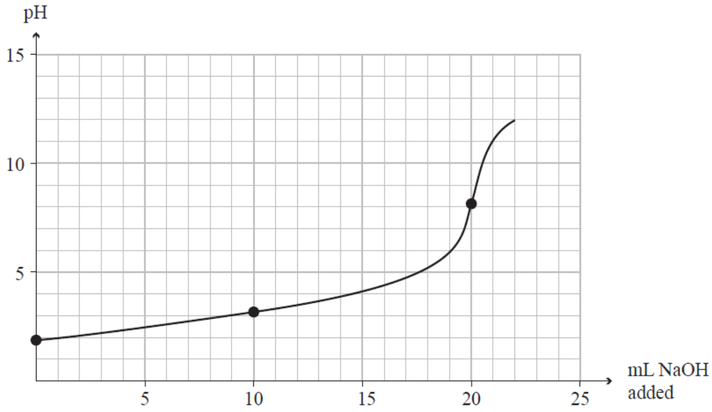
2014:3. (c) (i) Write the formulae of the four chemical species, apart from water and OH–, that are present at the point marked B on the curve (above).

(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

Past NCEA questions Titration Calculations (Part TWO)

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.

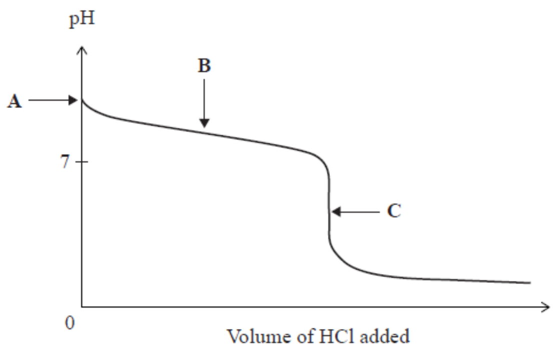
2015: 3a: 20.0 mL of 0.258 mol L^{-1} hydrofluoric acid, HF, solution is titrated with a sodium hydroxide, NaOH, solution. $HF + NaOH \rightarrow NaF + H_2O$ $pKa(HF) = 3.17$
(iii) After a certain volume of NaOH solution has been added, the concentration of HF in the solution will be twice that of the 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.
$pKa(CH_3COOH) = 4.76$ 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:
$NH_3 + HCl \rightarrow NH_4^+ + Cl^-$ $pK_a(NH_4^+) = 9.24$
The curve for this titration is given below.
Explain why the pH at the equivalence point (point C) is not 7.



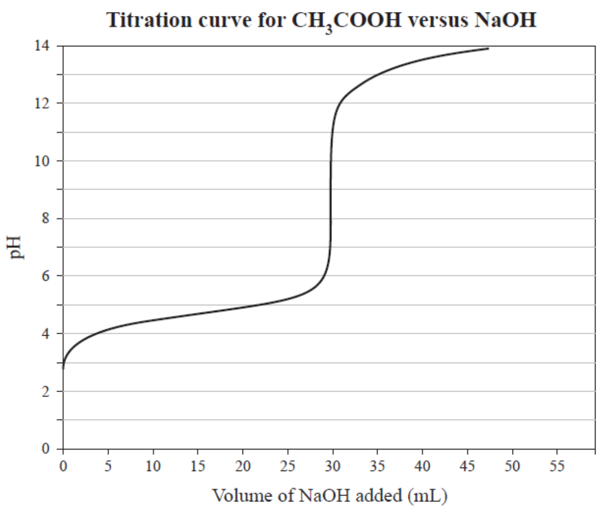
2016: 3b. Explain, in terms of the species present, why the pH at B (half way to the equivalence point) is 9.24.
3c. Show, by calculation, that the pH at the equivalence point (point C) is 4.96.

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

Past NCEA questions Titration Calculations (Part THREE)

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	

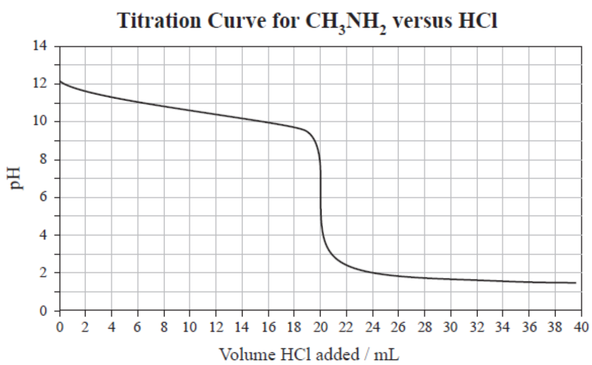
(b) (i) The ethanoic acid solution, $CH_3COOH_{(aq)}$, has a pH of 2.77 before any NaOH is added. Show by calculation that the concentration of the CH_3COOH solution is 0.166 mol L^{-1}.

(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.*

2018: 2. A titration was carried out by adding 0.210 mol L^{-1} hydrochloric acid, HCl, to 25.0 mL of 0.168 mol L^{-1} methanamine, CH_3NH_2.
The equation for the reaction is: $HCl + CH_3NH_2 \rightarrow CH_3NH_3^+ + Cl^-$
$pK_a(CH_3NH_3^+) = 10.6$ $K_a(CH_3NH_3^+) = 2.51 \times 10^{-11}$
(ii) Calculate the pH at the equivalence point.

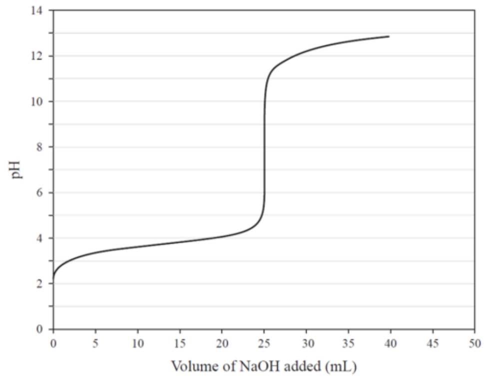


Past NCEA questions Titration Calculations (Part FOUR)

2019: Question: 2a A titration was carried out by adding 0.140 mol L^{-1} sodium hydroxide, NaOH, to 20.0 mL of 0.175 mol L^{-1} methanoic acid, HCOOH. The equation for the reaction is:

$HCOOH + NaOH \rightarrow HCOONa + H_2O$ $\quad$ $pK_a(HCOOH) = 3.74$

$K_a(HCOOH) = 1.82 \times 10^{-4}$



(i) With reference to the titration curve, put a tick next to the indicator most suited to identify the equivalence point. Explain your choice, including the consequences of choosing the other indicators.

Indicator	pK_a	**Tick ONE box below**
Thymol blue	1.70	
Bromocresol green	4.70	
Cresol red	8.30	

(ii) After 12.5 mL of NaOH has been added, the solution has a pH of 3.74. Explain the significance of this pH with reference to the relative concentrations of the species present. *No calculations are necessary.*

(ii) Calculate the pH at the equivalence point. V(NaOH) = 25mL from graph

(iv) Calculate the pH of the solution after 28.0 mL of 0.140 mol L^{-1} NaOH has been added.

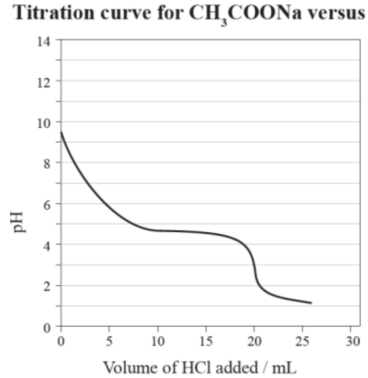
2020: Question 3a: A titration was carried out by adding 0.280 mol L^{-1} hydrochloric acid, HCl, to 25.0 mL of 0.224 mol L^{-1} sodium ethanoate solution, CH_3COONa.

The equation for the reaction is:

$CH_3COONa + HCl \rightarrow CH_3COOH + NaCl$ $\quad$ $pK_a\ (CH_3COOH) = 4.76$

$K_a(CH_3COOH) = 1.74 \times 10^{-5}$

Titration curve for CH_3COONa versus HCl



(i) List all the species present in a solution of sodium ethanoate. Do not include water

(ii) Calculate the pH of the 0.224 mol L–1 sodium ethanoate solution before any hydrochloric acid is added.

Question 3c: (i) Calculate the pH at the equivalence point.

Question 3c: (ii) In a second titration, 25.0 mL of 0.224 mol L^{-1} methanamine, CH_3NH_2, is titrated with the same 0.280 mol L^{-1} hydrochloric acid. $K_a\ (CH_3NH_3^+) = 2.29 \times 10^{-11}$ $\quad$ $K_a\ (CH_3COOH) = 1.74 \times 10^{-5}$

For this second titration, circle how the pH at the equivalence point will compare to the pH at the equivalence point in the titration of sodium ethanoate. Explain your answer. No calculations are necessary.

Lower pH $\quad$ Same pH $\quad$ Higher pH

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

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 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 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] x [ion2]^2 *3sgf (has no units)*	Q = [ion1] x [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 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

<table>
<tr><th colspan="2">pH Calculations QUESTION 1. (4 steps excellence)</th></tr>
<tr><td colspan="2">Question: Calculate the pH of a 0.109 mol L^{-1} solution of ethanamine.
pK_a($CH_3CH_2NH_3^+$) = 10.6
K_w = 1.00×10^{-14}</td></tr>
<tr><th colspan="2">ANSWER</th></tr>
<tr><td>1. determine if the solution is acid or base (will it accept or donate H^+) – strong or weak

And write down all available information</td><td>Ethanamine is a weak base
c(ethanamine) = 0.109 mol L^{-1}
pK_a($CH_3CH_2NH_3^+$) = 10.6
K_w = 1.00×10^{-14}</td></tr>
<tr><td>2. convert pK_a to K_a

$K_a = 10^{-pKa}$</td><td>$K_a = 10^{-pKa}$
$K_a = 10^{-10.6}$
$K_a = 2.51 \times 10^{-11}$</td></tr>
<tr><td>3. calculate [H_3O^+]
$[H_3O^+] = \sqrt{\frac{Ka \times Kw}{[base]}}$

3sgf and units</td><td>$[H_3O^+] = \sqrt{\frac{Ka \times Kw}{[base]}}$

$[H_3O^+] = \sqrt{\frac{2.51 \times 10^{-11} \times 1.00 \times 10^{-14}}{0.109 \text{ mol L}^{-1}}}$

$[H_3O^+] = 1.52 \times 10^{-12}$ molL^{-1}</td></tr>
<tr><td>4. calculate pH
pH = -log [H_3O^+]

3sgf
Double check answer against expected pH for your solution</td><td>pH = -log [H_3O^+]
pH = -log [1.52×10^{-12} molL^{-1}]
pH = 11.8

(pH range for weak base is 8-12) yes</td></tr>
<tr><th colspan="2">pH Calculations QUESTION 2. (3 steps Merit)</th></tr>
<tr><td colspan="2">Question: Calculate the pH of 0.0152 mol L^{-1} CH_3NH_3Cl solution.
K_a($CH_3NH_3^+$) = 2.29×10^{-11}</td></tr>
<tr><th colspan="2">ANSWER</th></tr>
<tr><td>1. determine if the solution is acid or base (will it accept or donate H^+) – strong or weak

And write down all available information</td><td>CH_3NH_3Cl is a weak acid (salt)
c(CH_3NH_3Cl) = 0.0152 mol L^{-1}
K_a($CH_3NH_3^+$) = 2.29×10^{-11}</td></tr>
<tr><td>2. calculate [H_3O^+]
$[H_3O^+] = \sqrt{Ka \times c(HA)}$

3sgf and units</td><td>$[H_3O^+] = \sqrt{Ka \times c(HA)}$
$[H_3O^+] = \sqrt{2.29 \times 10^{-11} \times 0.0152 \text{ mol L}^{-1}}$
$[H_3O^+] = 5.90 \times 10^{-7}$ mol L^{-1}</td></tr>
<tr><td>3. calculate pH
pH = -log [H_3O^+]

3sgf
Double check answer against expected pH for your solution</td><td>pH = -log [H_3O^+]
pH = -log [5.90×10^{-7} mol L^{-1}]
pH = 6.23

(pH range for weak acid is 3-6.9) yes</td></tr>
</table>

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.

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	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] = $[H_3O^+]$* *$[OH^-] = K_w / [H_3O^+]$* $[CH_3NH_2] = \frac{2.29 \times 10^{-11} \times 1.00 \times 10^{-14}}{(1.58 \times 10^{-12} \text{ mol}L^{-1})^2}$ $[CH_3NH_2]$ = 0.0912 mol L^{-1}

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

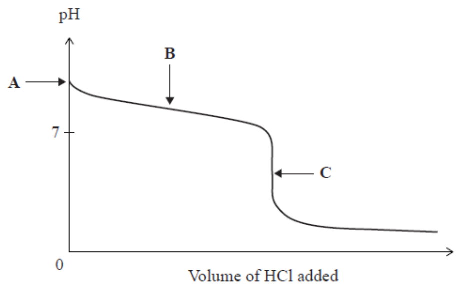
Titration Curve – after the Start pH QUESTION	
Question: : 20.0 mL of 0.0896 mol L^{-1} ethanoic acid is titrated with 0.100 mol L^{-1} sodium hydroxide. p*K*a (CH_3COOH) = 4.76 Calculate the pH of the titration mixture after 5.00 mL of NaOH has been added. $K_w = 1 \times 10^{-14}$	
ANSWER	
1. determine 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 $n = 0.0896 \times (20 \times 10^{-3})$ $n = 1.79 \times 10^{-3}$ mol
4. calculate *n*(NaOH) and therefore n(CH_3COO-) n = cv *assume n*(NaOH) = *n*(CH_3COO-) *3sgf and units*	n = cv $n = 0.1 \times (5 \times 10^{-3})$ $n = 5.00 \times 10^{-4}$ 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 $pK_a = 4.76$ $pH = pK_a + \log \frac{[CH_3COO\text{-}]}{[CH_3COOH]}$ *3sgf* *Check pH against estimate on curve*	$pH = pK_a + \log \frac{[CH_3COO\text{-}]}{[CH_3COOH]}$ $pH = 4.75 + \log \frac{[0.0200]}{[0.0515]}$ pH = 4.35

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 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.
$pK_a(NH_4^+) = 9.24$
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} $pK_a(NH_4^+) = 9.24$
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 $n(NH_3) = 6.40 \times 10^{-3}$ 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* $B^+ = HA$	$[H_3O^+] = \sqrt{Ka \times c(B^+)}$ $[H_3O^+] = \sqrt{10^{-9.24} \times 0.213\ molL^{-1}}$ $[H_3O^+] = 1.11 \times 10^{-5}\ molL^{-1}$
7. Calculate pH pH = -log $[H_3O^+]$ *3sgf* *Check pH against estimate on curve*	pH = -log $[H_3O^+]$ pH = -log $[1.11 \times 10^{-5}\ molL^{-1}]$ 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.