2017
Version

Chemistry AS 91392

C3.6 Aqueous Systems

NCEA Revision 2013 - 2016

GZ Science Resources

NCEA 2013 solubility

Question: 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 $Ag_2CrO_{4(s)}$.

NCEA 2013 solubility

Achieved Question

Question: 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 $Ag_2CrO_{4(s)}$.

$$K_s = [Ag^+]^2[CrO_4^{2-}]$$

GZ Science Resources

NCEA 2013 solubility

Question: 2b: 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}

(i) Calculate the solubility of $Ag_2CrO_{4(s)}$, and hence give the $[Ag^+]$ and $[CrO_4^{2-}]$ in the solution.

(ii) Determine the K_s(Ag2CrO4).

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NCEA 2013 solubility

Excellence Question

Question: 2b: 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}

(i) Calculate the solubility of $Ag_2CrO_{4(s)}$, and hence give the $[Ag^+]$ and $[CrO_4^{2-}]$ in the solution.

$$K_s = [Ag^+]^2[CrO_4^{2-}]$$

$$n(Ag_2CrO_4) = \frac{1.44 \times 10^{-3}}{332}$$

$$= 4.33 \times 10^{-6} \text{ mol in 50 mL}$$

$$[Ag_2CrO_4] = \frac{4.33 \times 10^{-6}}{50 \times 10^{-3}}$$

$$= 8.67 \times 10^{-5} \text{ mol L}^{-1}$$

$$[Ag^+] = 8.67 \times 10^{-5} \times 2 = 1.73 \times 10^{-4} \text{ mol L}^{-1}$$

$$[CrO_4^{2-}] = 8.67 \times 10^{-5} \text{ mol L}^{-1}$$

(ii) Determine the *K*s(Ag2CrO4).

$$K_s = (1.73 \times 10^{-4})^2 (8.67 \times 10^{-5})$$

$$= 2.61 \times 10^{-12}$$

NCEA 2014 solubility

Question: 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 $PbCl_2$.

Question: 2a: (ii) Write the expression for $K_s(PbCl_2)$.

NCEA 2014 solubility

Achieved Question

Question: 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 $PbCl_2$.

$$PbCl_{2(s)} \rightleftharpoons Pb^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)}$$

Question: 2a: (ii) Write the expression for $K_s(PbCl_2)$.

$$K_s = [Pb^{2+}][Cl^{-}]^2$$

NCEA 2014 solubility

Question: 2a: (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

NCEA 2014 solubility

Merit Question

Question: 2a: (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

$[Pb^{2+}] = x$ $[Cl^{-}] = 2x$

$K_s = 4x^3$

$$x = \sqrt[3]{\frac{K_s}{4}}$$

$$= \sqrt[3]{\frac{1.70 ´ 10^{-5}}{4}}$$

$$= 1.62 ´ 10^{-2} \text{ mol L}^{-1}$$

$[Pb^{2+}] = 1.62 \times 10^{-2}$ mol L^{-1}

$[Cl^{-}] = 3.24 \times 10^{-}$ mol L^{-1}

$$s = \frac{\sqrt[3]{K_s}}{4}$$

NCEA 2015 Solubility

Question: 2a: (i) Sufficient calcium carbonate, $CaCO_{3(s)}$, is dissolved in water to make a saturated solution.

- Write the equation for the equilibrium occurring in a saturated solution of $CaCO_3$.

Question: 2a: (ii) Write the expression for Ks($CaCO_3$).

NCEA 2015 Solubility

Achieved Question

Question: 2a: (i) Sufficient calcium carbonate, $CaCO_{3(s)}$, is dissolved in water to make a saturated solution.

- ❑ Write the equation for the equilibrium occurring in a saturated solution of $CaCO_3$.

$$CaCO_3(s) \rightleftharpoons Ca^{2+}(aq) + CO_3^{2-}(aq)$$

Question: 2a: (ii) Write the expression for *K*s($CaCO_3$).

$$K_s = [Ca^{2+}][CO_3^{2-}]$$

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NCEA 2015 Solubility

Question: 2a: (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}.

Photo: http://www.chemistryland.com

Merit Question

Question: 2a: (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}.

$$\mathbf{K_s(CaCO_3) = (5.74 \times 10^{-5})^2 = 3.29 \times 10^{-9}}$$

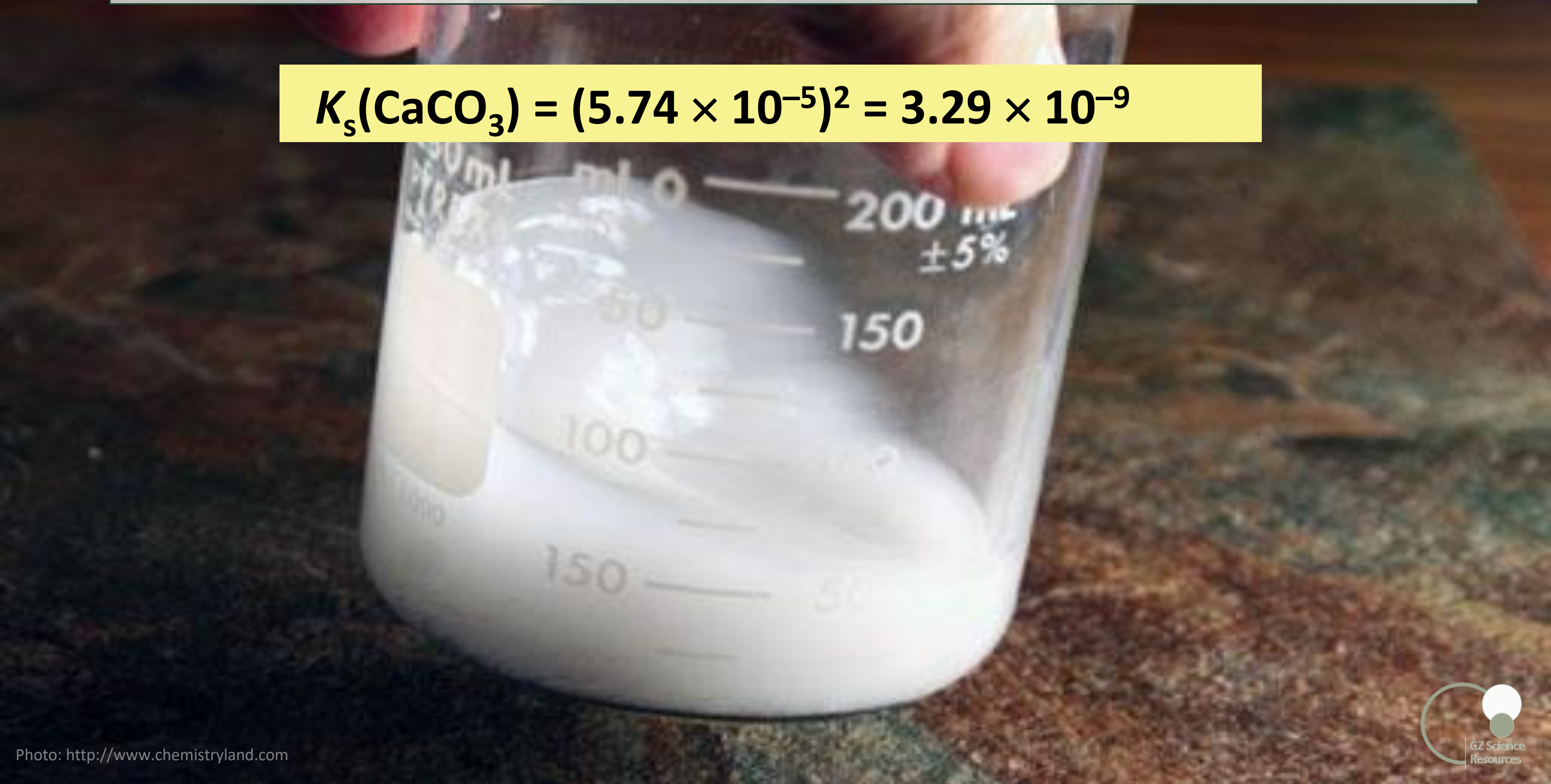


Photo: http://www.chemistryland.com

NCEA 2016 Solubility

Question: 1a: Silver carbonate, Ag_2CO_3, is a sparingly soluble salt.

$K_s(Ag_2CO_3) = 8.10 \times 10^{-12}$ at 25ºC $M(Ag_2CO_3) = 276 \text{ g mol}^{-1}$

(a) Write the solubility product expression, K_s, for silver carbonate (Ag_2CO_3).

NCEA 2016 Solubility

Achieved Question

Question: 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}

(a) Write the solubility product expression, K_s, for silver carbonate (Ag_2CO_3).

$$K_s = [Ag^+]^2[CO_3^{2-}]$$

NCEA 2016 Solubility

Question: 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.

NCEA 2016 Solubility

Excellence Question

Question: 1b: Silver carbonate, Ag_2CO_3, is a sparingly soluble salt.

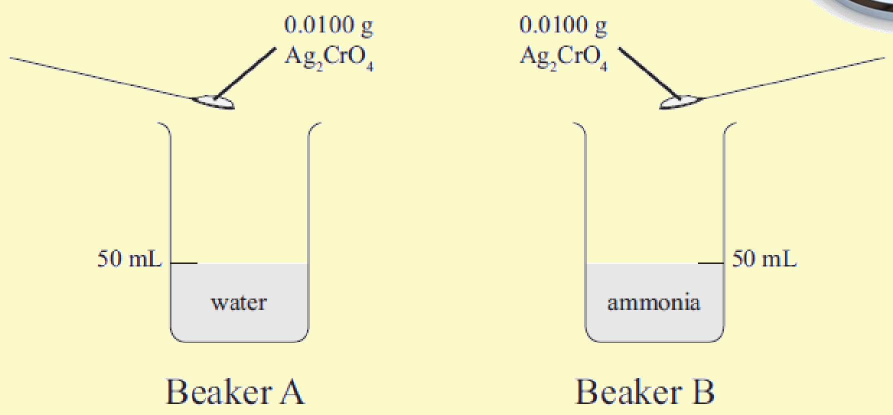
$K_s(Ag_2CO_3) = 8.10 \times 10^{-12}$ at 25ºC $M(Ag_2CO_3) = 276 \text{ g mol}^{-1}$

Calculate the mass of Ag_2CO_3 that will dissolve in 50 mL of water to make a saturated solution at 25ºC.

Let s = solubility $[Ag^+] = 2s$ $[CO_3^{2-}] = s$ $K_s = 4s^3$ $s = 1.27 \times 10^{-4}$ mol L^{-1} $n = c \times v$ = 6.33×10^{-6} mol $m = n \times M$ = 1.75×10^{-3} g	**OR** $g\ L^{-1} = c \times M$ = 0.0349 g L^{-1} so mass in 50 mL = $\frac{0.0349 \times 50}{1000}$ $= 1.75 \times 1^{-3}$ g

NCEA 2013 solubility and base

Question: 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.



NCEA 2013 solubility and base

Excellence Question

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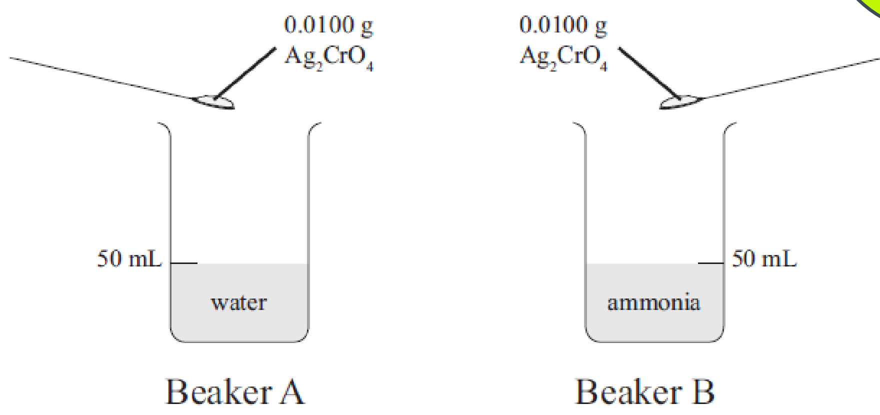


Dissolving 0.0100g of silver chromate in 50 mL water will result in solid being present, as the required amount to make a saturated solution is 1.44×10^{-3} g in 50 mL, so any more than this will form a solid.

If the same mass is added to 50 mL of ammonia, more will dissolve and less solid will be present due to the formation of a complex ion.

The Ag_2CrO_4 will dissociate completely and form an equilibrium.

$Ag_2CrO_4 \rightleftharpoons 2Ag^+ + CrO_4^{2-}$

$Ag^+ + 2NH_3 \rightarrow [Ag(NH_3)_2]^+$

The silver ion will then react further with NH_3, removing it from the above equilibrium. Thus, more Ag_2CrO_4 will dissolve to re-establish equilibrium.

NCEA 2014 solubility and acid

Question: 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.

NCEA 2014 solubility and acid

Excellence Question

Question: 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.

$Zn(OH)_2(s) \rightleftharpoons Zn^{2+}(aq) + 2OH^-(aq)$ When pH is less than 4 / low, $[OH^-]$ is decreased due to the reaction with H_3O^+ to form water,

$H_3O^+ + OH^- \rightarrow H_2O$ so equilibrium shifts to the right to produce more $[OH^-]$, therefore more $Zn(OH)_2$ will dissolve.

When pH is greater than 10 / high, then more OH^- is available and the complex ion (zincate ion) will form.

$Zn(OH)_2(s) + 2OH^- \rightarrow [Zn(OH)_4]^{2-}$ OR $Zn^{2+} + 4OH^- \rightarrow [Zn(OH)_4]^{2-}$

This decrease in $[Zn^{2+}]$ causes the position of equilibrium to shift further to the right, therefore more $Zn(OH)_2$ dissolves.

NCEA 2015 Solubility and Acid

Question: 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.

NCEA 2015 Solubility and Acid

Merit Question

Question: 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.

The H_3O^+ from the acidic solution reacts with the CO_3^{2-}. This reduces $[CO_3^{2-}]$, causing the equilibrium to shift towards the products / RHS to replace some of the lost CO_3^{2-}. Therefore more solid $CaCO_3$ will dissolve.

$$2H_3O^+ + CO_3^{2-} \rightarrow 3H_2O + CO_2$$

NCEA 2016 Solubility and Acid

Question: 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.

NCEA 2016 Solubility and Acid

Excellence Question

Question: 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.

$$Ag_2CO_{3(s)} \rightarrow 2Ag^+_{(aq)} + CO_3^{2-}{}_{(aq)}$$

$$Ag^+_{(aq)} + 2NH_{3(aq)} \leftrightharpoons [Ag(NH_3)_2]_{(aq)}$$

Then when a base is added to this system then it will react with the cation effectively reducing the concentration of this in the solution.

The equilibrium responds by favouring the forward reaction and thus more dissolves.

Base "locks up" many cations into complex ions

NCEA 2014 common ion effect

Question: 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}

NCEA 2014 common ion effect

Excellence Question

Question: 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}

$$n(Pb(NO_3)_2) = \frac{2.00\text{ g}}{331\text{ g mol}^{-1}}$$

$= 6.04 \times 10^{-3}$ mol

$\therefore [Pb^{2+}] = 6.04 \times 10^{-3}$ mol / 0.500L

$= 1.21 \times 10^{-2}$ mol L^{-1}

$Q = (1.21 \times 10^{-2}) \text{ x } (0.440)^2$

$= 2.34 \times 10^{-3}$

As $Q > K_s$, **a precipitate will form.**

NCEA 2015 Common Ion effect

Question: 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

The ratio of the concentrations of products and reactants is called Q.

NCEA 2015 Common Ion effect

Excellence Question

Question: 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

The ratio of the concentrations of products and reactants is called Q.

$Pb(OH)_2 \rightleftharpoons Pb^{2+} + 2OH^-$

$Q = [Pb^{2+}][OH^-]^2$
$[Pb^{2+}] = 0.5 \times 0.00421 = 2.105 \times 10^{-3}$
$[OH^-] = 0.5 \times 0.0398 = 1.99 \times 10^{-2}$
$Q = (2.105 \times 10^{-3}) \times (1.99 \times 10^{-2})^2$
$Q = 8.34 \times 10^{-7}$
Since $Q > K_s$, a precipitate of $Pb(OH)_2$ will form.

pH =12.6
pOH = 1.4
$[OH^-] = 0.0398$

NCEA 2016 Common Ion effect

Question: 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

The ratio of the concentrations of products and reactants is called Q.

NCEA 2016 Common Ion effect

Excellence Question

Question: 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

$AgNO_3$ dilution: $\frac{20}{55} \times 0.105$ = *0.0382*

Na_2CO_3 dilution: $\frac{35}{55} \times 0.221$ = *0.141*

Q / I.P. = $[0.03818]^2[0.1406] = 2.06 \times 10^{-4}$

As Q / I.P. > K_s, a precipitate will form.

The ratio of the concentrations of products and reactants is called Q.

$$K_s = [Ag^+]^2[CO_3^{2-}]$$

2015 dissociation equations - NCEA Case Study

Question: 1a: (i) Methylammonium chloride, CH_3NH_3Cl, dissolves in water to form a weakly acidic solution.

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

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

2014 dissociation equations - NCEA Case Study

Question: 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 dissociation equations - NCEA Case Study

Achieved Question

Question: 1a: (i) Methylammonium chloride, CH_3NH_3Cl, dissolves in water to form a weakly acidic solution.

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

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

$CH_3NH_3Cl(s) \rightarrow CH_3NH_3^+(aq) + Cl^-(aq)$ $CH_3NH_3^+ + H_2O \rightleftharpoons CH_3NH_2 + H_3O^+$

2014 dissociation equations - NCEA Case Study

Question: 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.

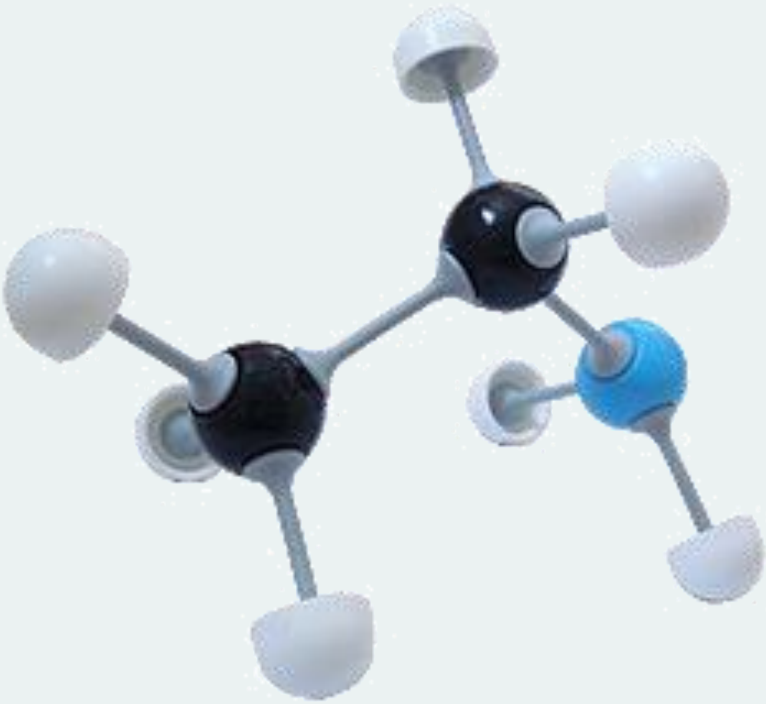
$$\mathbf{HOCl + H_2O \rightleftharpoons OCl^- + H_3O^+}$$

NCEA 2016 dissociation equations

Question: 2a: (i) Ethanamine, $CH_3CH_2NH_2$, is a weak base.

$pK_a(CH_3CH_2NH_3^+) = 10.6$ $K_a(CH_3CH_2NH_3^+) = 2.51 \times 10^{-11}$

(a) Write an equation to show the reaction of ethanamine with water.



NCEA 2016 dissociation equations

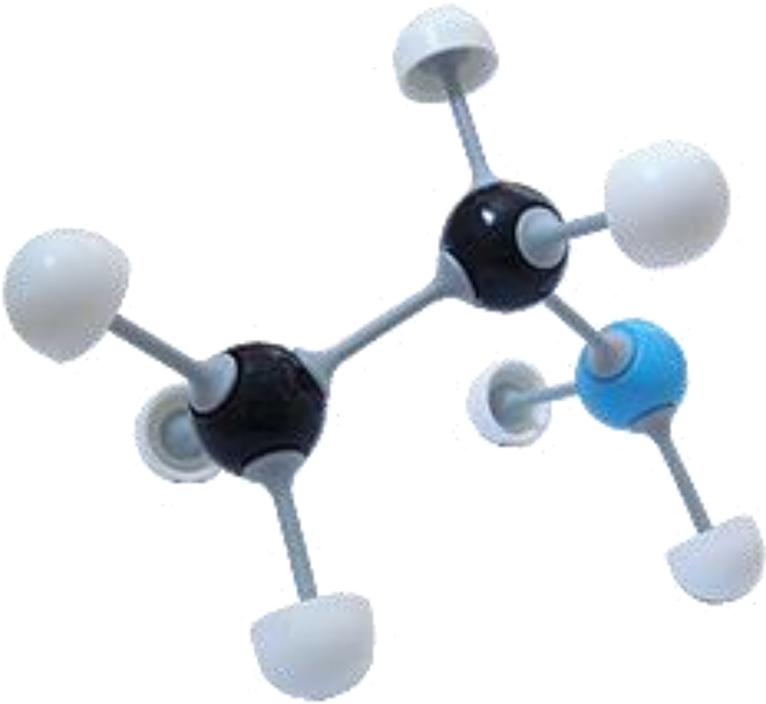
Achieved Question

Question: 2a: (i) Ethanamine, $CH_3CH_2NH_2$, is a weak base.

p$K_a(CH_3CH_2NH_3^+) = 10.6$ $K_a(CH_3CH_2NH_3^+) = 2.51 \times 10^{-11}$

(a) Write an equation to show the reaction of ethanamine with water.

$$CH_3CH_2NH_2 + H_2O \rightleftharpoons CH_3CH_2NH_3^+ + OH^-$$



NCEA 2013 Species present (pH)

Question: 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

NCEA 2013 Species present (pH)

Excellence Question

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CH_3NH_3Cl CH_3NH_2 HCl

$HCl < CH_3NH_3Cl < CH_3NH_2$

- HCl, a strong acid, reacts completely with water to form 1 mol L^{-1} H_3O^+ and hence a low pH. $HCl + H_2O \rightarrow H_3O^+ + Cl^-$
- CH_3NH_3Cl dissociates completely in water to form $CH_3NH_3^+$ and Cl^-. $CH_3NH_3^+$, a weak acid, partially reacts with water to form less than 1 mol L^{-1} H_3O^+ and hence a higher pH than HCl. $CH_3NH_3Cl \rightarrow CH_3NH_3^+ + Cl^-$

  $CH_3NH_3^+ + H_2O \leftrightarrow CH_3NH_2 + H_3O^+$
- CH_3NH_2, a weak base, partially reacts with water to form OH^- ions. So there are more OH^- ions than H_3O^+ ions and the pH is thus high.

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

NCEA 2014 Species present

Question: 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.

GZ Science Resources

NCEA 2014 Species present

Excellence Question

Question: 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.

$HOCl > H_3O^+ > OCl^- > OH^-$ or $HOCl > H_3O^+ = OCl^- > OH^-$

HOCl partially dissociates, and so the equilibrium lies to the LHS/favours the reactants; therefore HOCl is present in the greatest amounts.

H_3O^+ and OCl^- are produced in equal amounts / there is a small contribution to H_3O^+ from water therefore $H_3O^+ > OCl^-$

Because there is a relatively high $[H_3O^+]$, the $[OH^-]$ is very low (or links to K_w).

NCEA 2016 Species Present

Question: 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.

NCEA 2016 Species Present

Excellence Question

Question: 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.

$Cl^- > CH_3NH_3^+ > H_3O^+ = CH_3NH_2 > OH^-$ OR

$Cl^- > CH_3NH_3^+ > H_3O^+ > CH_3NH_2 > OH^-$

$CH_3CH_2NH_3Cl \rightarrow CH_3CH_2NH_3^+ + Cl^-$

$CH_3CH_2NH_3Cl$ completely dissociates.

(The chloride ion does not react further with water and so will be in the greatest concentration.)

The ethanamine ion will react further with water, but only partially, leaving it the next in the series.

$CH_3CH_2NH_3^+ + H_2O \rightleftharpoons CH_3CH_2NH_2 + H_3O^+$

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).

OH^- is present in the lowest concentration as this comes from the dissociation of water only.

NCEA 2013 Conductivity

Question: 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.

NCEA 2013 Conductivity

Excellence Question

Question: 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.

HCl = CH_3NH_3Cl > CH_3NH_2

CH_3NH_3Cl and HCl will dissociate completely in water to produce 2 mol L^{-1} ions.

CH_3NH_2 will only partially react with water to produce less than 1 mol L^{-1} of ions.

NCEA 2015 Conductivity

Question: 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

NCEA 2015 Conductivity

Excellence Question

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

The pH of a solution is calculated from its $[H_3O^+]$.

- **NaOH** is an ionic solid that is a strong base and dissociates completely to produce a high OH^- concentration (low $[H_3O^+]$). Since $[OH^-]$ is high / $[H_3O^+]$ is low, the pH is high.

$NaOH \rightarrow Na^+ + OH^-$

NCEA 2015 Conductivity

Excellence Question

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

- ❑ **CH_3NH_2** is a weak base that partially reacts / dissociates / ionises with H_2O producing a lower concentration of OH^-, Therefore it has a lower pH than NaOH:

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

- ❑ The **CH_3COONa** is an ionic solid that dissociates completely in H_2O. The CH_3COO^- ion is a weak base that partially reacts / dissociates / ionises with H_2O producing a lower concentration of OH^-.

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

The pH is closer to 7, showing it is the weakest base. Therefore it has a lowest pH

NCEA 2015 Conductivity

Excellence Question

Question: 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

Electrical conductivity:

Electrical conductivity is determined by the concentration of ions.

- ❑ **NaOH** completely dissolves to produce a high concentration of Na^+ and OH^- ions in solution.

$NaOH \rightarrow Na^+ + OH^-$ Therefore it is a good conductor.

NCEA 2015 Conductivity

Excellence Question

Question: 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

Electrical conductivity:

- Since **CH_3NH_2** is a weak base, it only partially reacts with water to produce a low concentration of ions in solution so it is a poor electrical conductor.

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

- **CH_3COONa** is also an ionic solid. It dissolves completely to produce a high concentration of Na^+ and CH_3COO^- ions:

$CH_3COONa \rightarrow Na^+ + CH_3COO^-$

Therefore it is a good conductor.

NCEA 2014 pH calculations

Question: 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.

NCEA 2014 pH calculations

Merit Question

Question: 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.

Hydrofluoric acid is a stronger acid/more acidic/dissociates more because it has a smaller pK_a (larger K_a) than hypochlorous acid.

So HF will therefore have a higher $[H_3O^+]$. As $[H_3O^+]$ increases, the pH decreases, so HF will have a lower pH than HOCl.

(pH HF = 2.09, HOCl = 4.27)

→ larger pKa more reactants, the weaker the acid

NCEA 2015 pH calculations

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

Ka($CH_3NH_3^+$) = 2.29×10^{-11}

NCEA 2015 pH calculations

Merit Question

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

*K*a($CH_3NH_3^+$) = 2.29×10^{-11}

$$\left[H_3O^+\right] = \sqrt{K_a} \times HA$$

$$= 5.90 \times 10^{-7}$$

$$pH = 6.23$$

$K_a = 10^{-Pka}$

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

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

$[H_3O^+] = 5.90 \times 10^{-7}$ mol L^{-1}
pH = $-\log 5.90 \times 10^{-7} = 6.23$

NCEA 2016 pH calculations

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

pK_a($CH_3CH_2NH_3^+$) = 10.6

NCEA 2016 pH calculations

Excellence Question

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

$pK_a(CH_3CH_2NH_3^+) = 10.6$

$[H_3O^+] = \sqrt{(K_a \times K_w \div [CH_3CH_2NH_2])}$

$[H_3O^+] = \sqrt{(2.51\times10^{-11} \times 1.00\times10^{-14} \div 0.109)}$

$[H_3O^+] = 1.52\times10^{-12}$ $molL^{-1}$

$pH = -\log[H_3O^+] = 11.8$

$K_b = 1 \times 10^{-14} / K_a$

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

$[H_30^+] = 1 \times 10^{-14}/[OH^-]$

$pH = -\log[H_30^+]$

NCEA 2013 Titration Curve - (PART ONE)

Question: 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

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

NCEA 2013 Titration Curve - (PART ONE)

Excellence Question

Question: 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

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

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

- $$pH = pK_a + \log\frac{[base]}{[acid]}$$

Alternative equation

$$[H_3O^+] = \sqrt{1.74 \times 10^{-5} \times 0.0896} \text{ mol L}^{-1}$$

$$= 1.25 \times 10^{-3} \text{ mol L}^{-1}$$

$$pH = -\log[H_3O^+] = 2.90$$

NCEA 2013 Titration Curve - (PART TWO)

Question: 3b: Halfway to the equivalence point of the titration, the pH = p*K*a of the ethanoic acid.

Discuss the reason for this.

NCEA 2013 Titration Curve - (PART TWO)

Excellence Question

Question: 3b: Halfway to the equivalence point of the titration, the pH = p*K*a of the ethanoic acid.

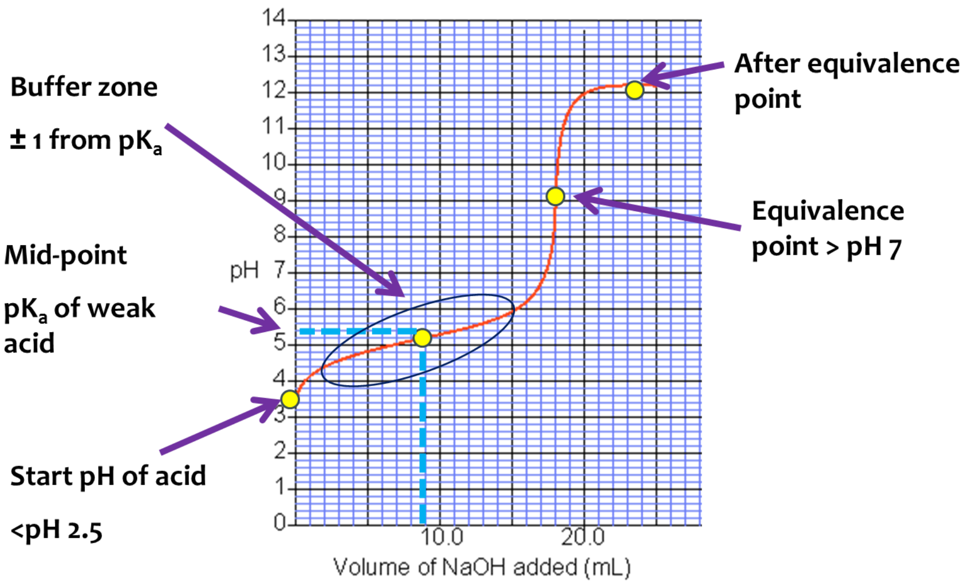
Discuss the reason for this.

Halfway to equivalence point, half of the ethanoic acid has been used up. There are now equimolar quantities of ethanoic acid and sodium ethanoate.

According to the equation when $[CH_3COOH] = [CH_3COO^-]$

then $Ka = [H_3O^+]$

So pKa = pH



NCEA 2013 Titration Curve - (PART THREE)

Question: 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.

NCEA 2013 Titration Curve - (PART THREE)

Excellence Question

Question: 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.

$NaOH(aq) + CH_3COOH(aq) \rightarrow NaCH_3COO(aq) + H_2O(l)$

- $[CH_3COO^-]$ increases as it is formed in reaction
- $[Na^+]$ increases as NaOH is added
- $[CH_3COOH]$ decreases as it reacts with NaOH
- $[H_3O^+]$ decreases because $[CH_3COO^-]$ / $[CH_3COOH]$ increases and K_a is a constant.
- $[OH^-]$ increases because $[H_3O^+]$ decreases and $[H_3O^+]$ $[OH^-]$ is constant.

GZ Science Resources

NCEA 2013 Titration Curve - (PART FOUR)

Question: 3c: 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

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

NCEA 2013 Titration Curve - (PART FOUR)

Excellence Question

Question: 3c: 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

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

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

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

$n(NaOH$ added$) = 0.1 \times (5 \times 10^{-3}) = 5 \times 10^{-4}$ mol **n = c x v**

After 5 mL NaOH added: (total 25mL)

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

($n(CH_3COOH$ – $n(NaOH)$ after 5mL)

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

$[CH_3COOH] = 0.0516$ mol L^{-1} **c = n / v**

$[CH_3COO^-] = 0.0200$ mol L^{-1} **c = n / v**

$[H_3O^+] = 4.48 \times 10^{-5}$ mol L^{-1}

pH = 4.35

Step Four: Calculate pH of the equivalence point (end point)

a) Use the number of moles (n) of base (as calculated in step two) required to completely react with n of acid present to reach equivalence.

Each 1 mole of base required to react with acid produces 1 mole of conjugate base.

Example

$HCOOH_{(aq)} + NaOH_{(aq)} \leftrightarrow HCOONa_{(aq)} + H_2O_{(l)}$

b) Calculate c(conjugate base) using c = n/v

V = initial volume in flask + volume added during titration to reach equivalence

c) Use pH equations

$$[H_3O^+] = \sqrt{\frac{k_a \times k_w}{c(\text{conjugate base})}}$$

NCEA 2014 Titration Curve - (PART ONE)

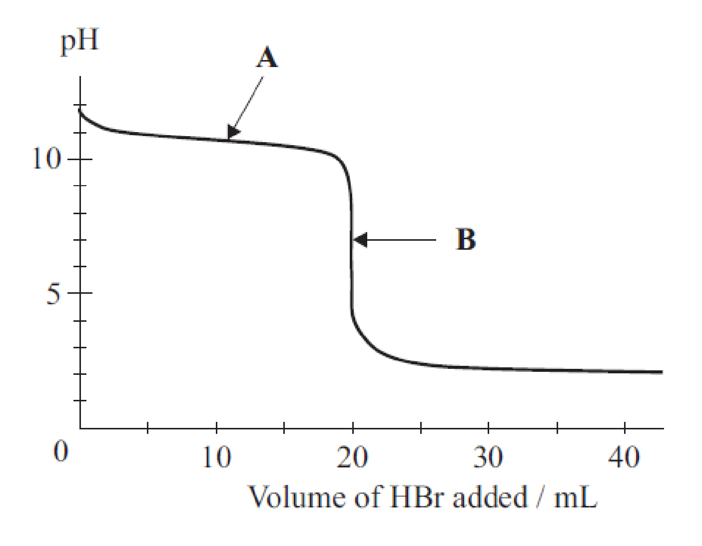
Question: 3a: A titration was carried out by adding hydrobromic acid, HBr, to 20.0 mL of aqueous methylamine, CH_3NH_2, solution.

The equation for the reaction is:

$CH_3NH_2 + HBr \rightarrow CH_3NH_3+ + Br^-$

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

Explain why the pH does not change significantly between the addition of 5 to 15 mL of HBr (around point **A** on the curve).



NCEA 2014 Titration Curve - (PART ONE)

Excellence Question

Question: 3a: 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:

$CH_3NH_2 + HBr \rightarrow CH_3NH_3+ + Br^-$

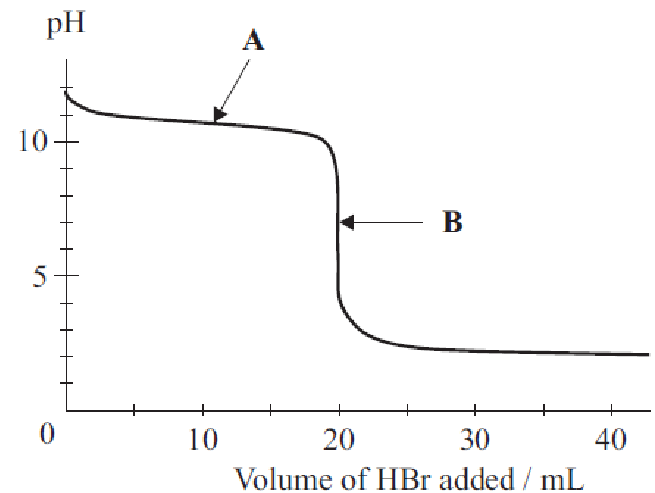
$K_a(CH_3NH_3^+) = 2.29 \times 10^{-11}$

Explain why the pH does not change significantly between the addition of 5 to 15 mL of HBr (around point **A** on the curve).

At point A, $[CH_3NH_2] \approx [CH_3NH_3^+]$. So the solution has buffering properties in the proximity of point A. When HBr is added, the H_3O^+ is consumed:

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

Since the H_3O^+ is removed from the solution (neutralised), the pH does not change significantly.



NCEA 2014 Titration Curve - (PART TWO)

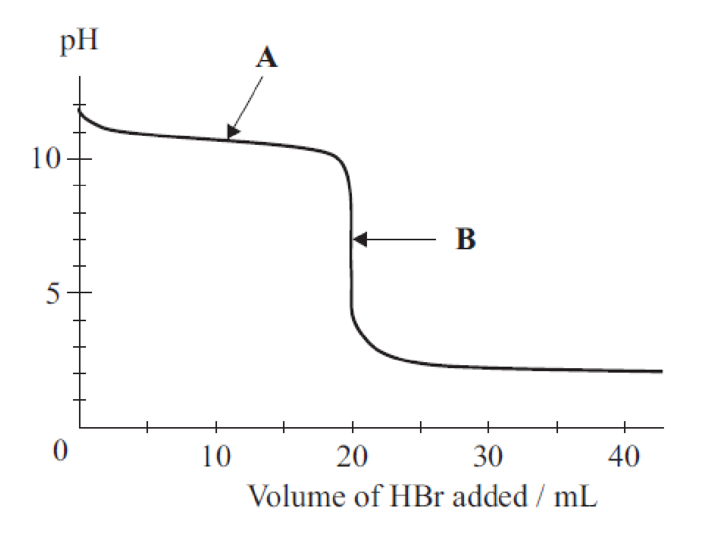
Question: 3b: A titration was carried out by adding hydrobromic acid, HBr, to 20.0 mL of aqueous methylamine, CH_3NH_2, solution.

The equation for the reaction is: $CH_3NH_2 + HBr \rightarrow CH_3NH_3+ + Br^-$

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

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 $L^{-1.}$



NCEA 2014 Titration Curve - (PART TWO)

Excellence Question

Question: 3b: 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: $CH_3NH_2 + HBr \rightarrow CH_3NH_3+ + Br^-$

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

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}.

$[H_3O^+] = 10^{-11.8} = 1.58 \times 10^{-12}$

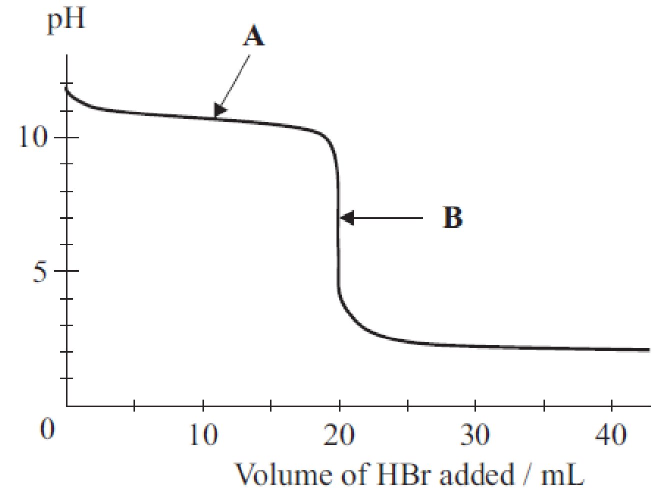
$$K_a = \frac{[CH_3NH_2][H_3O^+]}{[CH_3NH_3^+]}$$

$$= \frac{[CH_3NH_2][H_3O^+]}{[OH^-]}$$

$$2.29 \times 10^{-11} = \frac{[CH_3NH_2] \times (10^{-11.8})^2}{1 \times 10^{-14}}$$

$$[CH_3NH_2] = \frac{(2.29 \times 10^{-11}) \times (1 \times 10^{-14})}{(10^{-11.8})^2}$$

$= 0.0912$ mol L^{-1}



GZ Science Resources

NCEA 2014 Titration Curve - (PART THREE)

Question: 3c: Write the formulae of the four chemical species, apart from water and OH–, that are present at the point marked **B** on the curve.

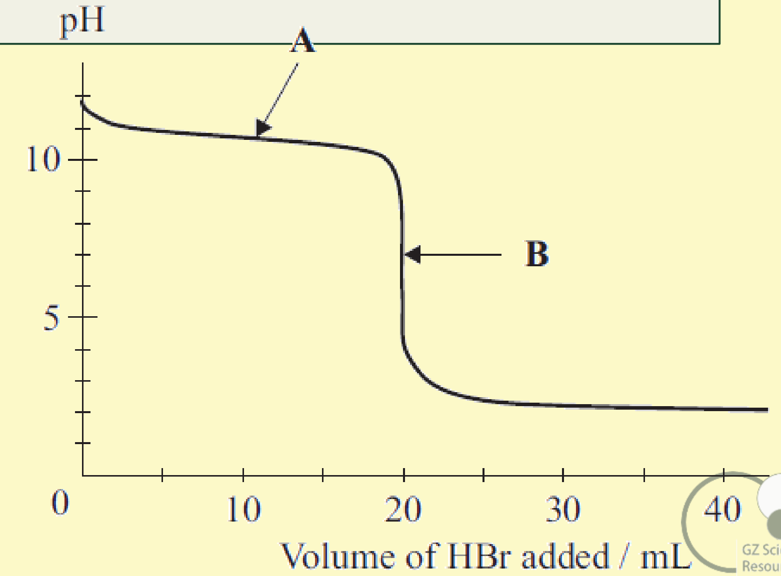
(ii) Compare and contrast the solution at point **B** with the initial aqueous methylamine solution.

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.



NCEA 2014 Titration Curve - (PART THREE)

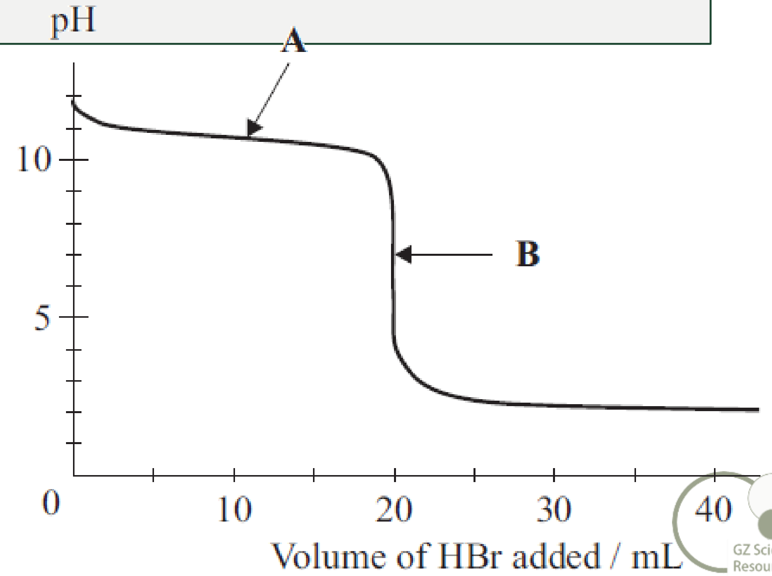
Excellence Question

Question: 3c: Write the formulae of the four chemical species, apart from water and OH–, that are present at the point marked **B** on the curve.

(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.



$CH_3NH_3^+$, Br^-, CH_3NH_2, H_3O^+

GZ Science Resources

NCEA 2014 Titration Curve - (PART FOUR)

Question: 3c: (ii) Compare and contrast the solution at point **B** with the initial aqueous methylamine solution.

NCEA 2014 Titration Curve - (PART FOUR)

Excellence Question

Question: 3c: (ii) Compare and contrast the solution at point **B** with the initial aqueous methylamine solution.

At the start, before addition of HBr there is a solution of weak base (CH_3NH_2) which only partially reacts with water to produce a relatively low concentration of ions.

As a result, the initial CH_3NH_2 solution will be a poor electrical conductor.

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

Therefore species present are $CH_3NH_2 > OH^- \geq CH_3NH_3^+ > H_3O^+$

At point B, there is a solution of the salt CH_3NH_3Br present which is dissociated completely into ions. Therefore there is a relatively high concentration of ions ($CH_3NH_3^+$ and Br^-) present in the solution, so it will be a good electrical conductor / electrolyte.

$CH_3NH_3Br \rightarrow CH_3NH_3^+ + Br^-$

$CH_3NH_3^+$ reacts with water according to the equation

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

Species present are $Br^- > CH_3NH_3^+ > H_3O^+ \geq CH_3NH_2 > (OH^-)$

GZ Science Resources

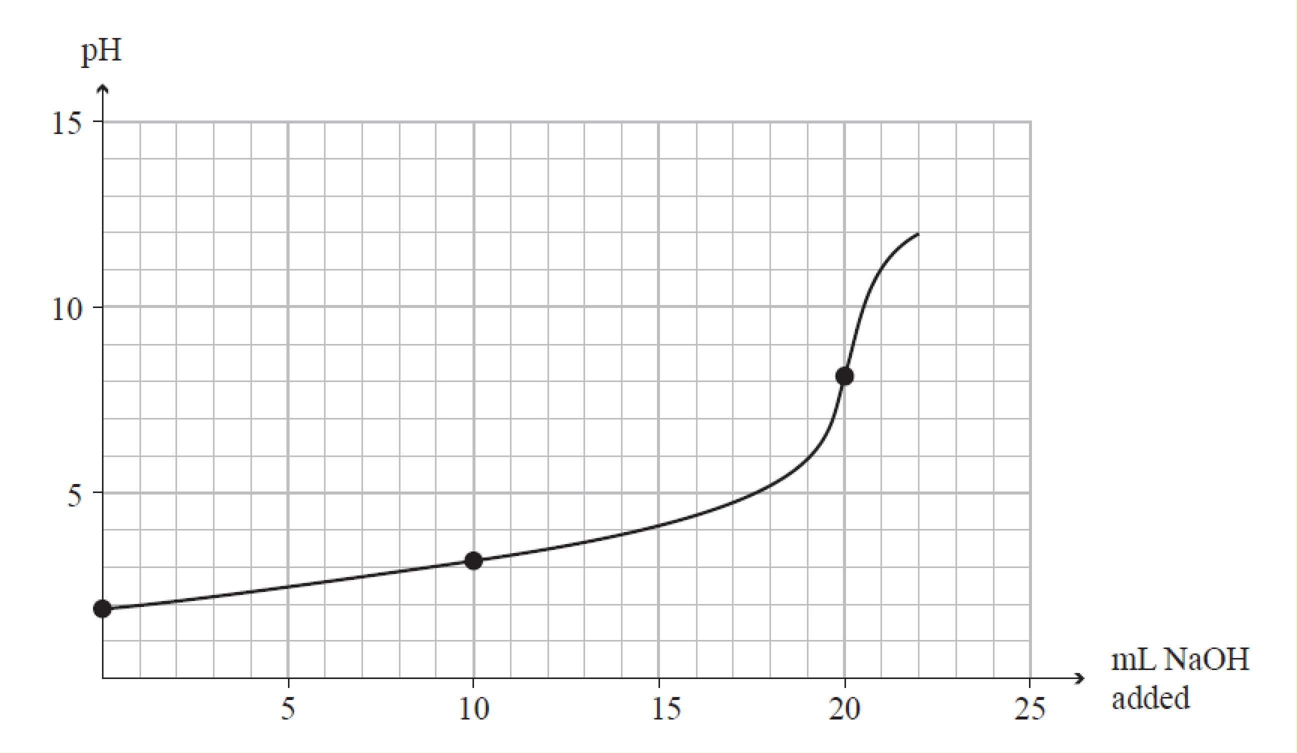
NCEA 2015 Titration Curves - (PART ONE)

Question: 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:

$HF + NaOH \rightarrow NaF + H_2O$ p*K*a(HF) = 3.17

(i) Identify the species in solution at the equivalence point.



NCEA 2015 Titration Curves - (PART ONE)

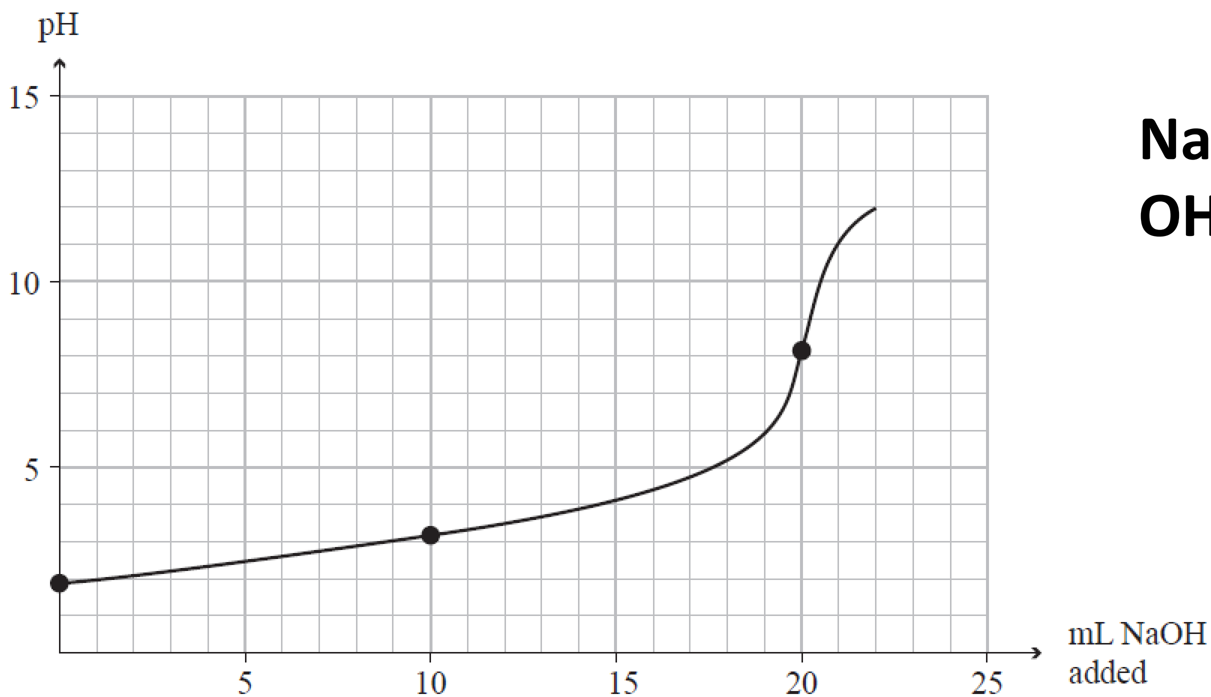
Achieved Question

Question: 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:

$HF + NaOH \rightarrow NaF + H_2O$ $\quad\quad$ p*K*a(HF) = 3.17

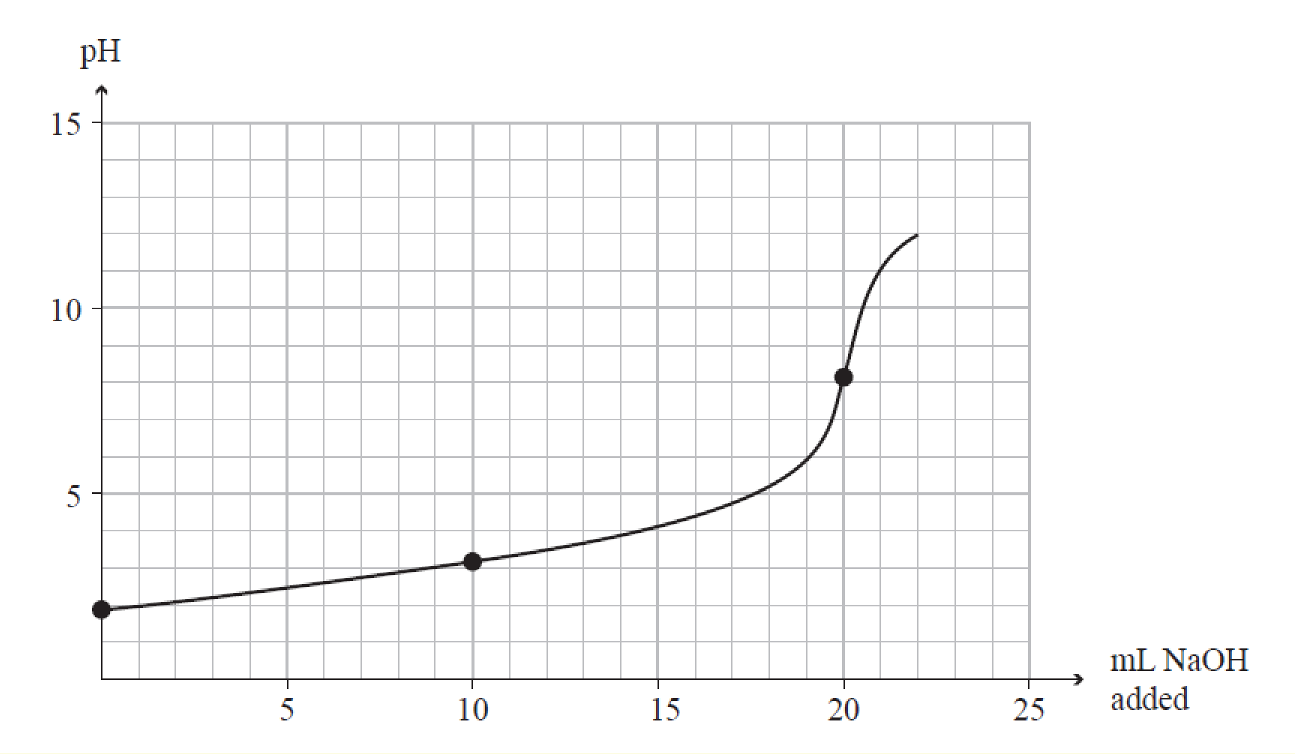
(i) Identify the species in solution at the equivalence point.



Na^+, F^-, H_2O, HF, OH^-, H_3O^+.

NCEA 2015 Titration Curves - (PART TWO)

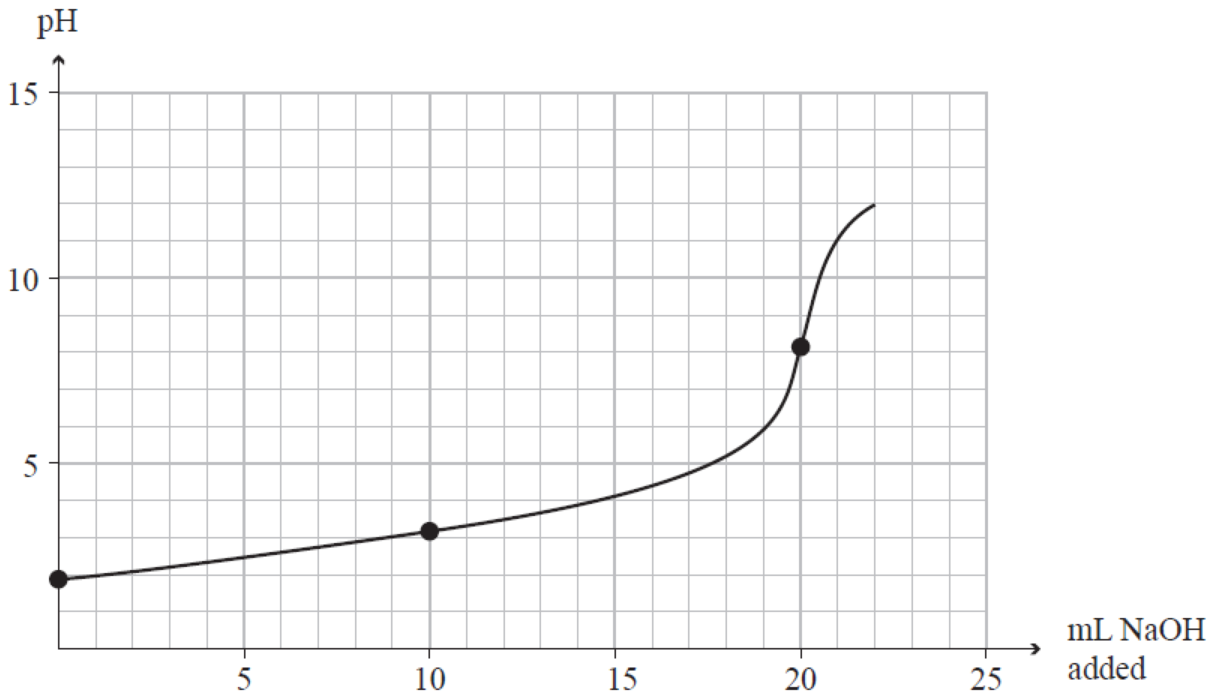
Question: 3a: (ii) Explain why the pH at the equivalence point is greater than 7. Include an equation in your answer.



NCEA 2015 Titration Curves - (PART TWO)

Merit Question

Question: 3a: (ii) Explain why the pH at the equivalence point is greater than 7. Include an equation in your answer.



A weak base, F^-, is present at the equivalence point:

$F^- + H_2O \rightleftharpoons HF + OH^-$

This increase in $[OH^-]$ causes the pH to be greater than 7.

NCEA 2015 Titration Curves - (PART THREE)

Question: 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.

NCEA 2015 Titration Curves - (PART THREE)

Achieved Question

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

$HF + NaOH \rightarrow NaF + H_2O$ $pK_a(HF) = 3.17$

(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.

$[H_3O^+] = 2 \times 10^{-3.17} = 1.35 \times 10^{-3}$ mol L^{-1}

$pH = -\log(1.35 \times 10^{-3}) = 2.87$.

$pH = pK_a + \log [F^-] / [HF]$

$= 3.17 + \log 0.5$

$= 2.87$

Alternative method

Since there are significant concentrations of the weak acid and its conjugate base the solution can resist added acid or base.

However, since the pH of the buffer solution is less than the pKa, / [HF] > [F–], it is more effective against added base than acid.

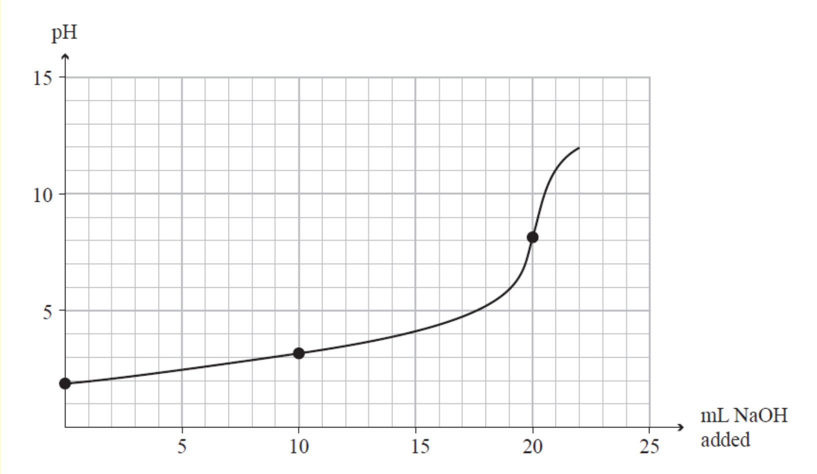
GZ Science Resources

NCEA 2015 Titration Curves - (PART FOUR)

Question: 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

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



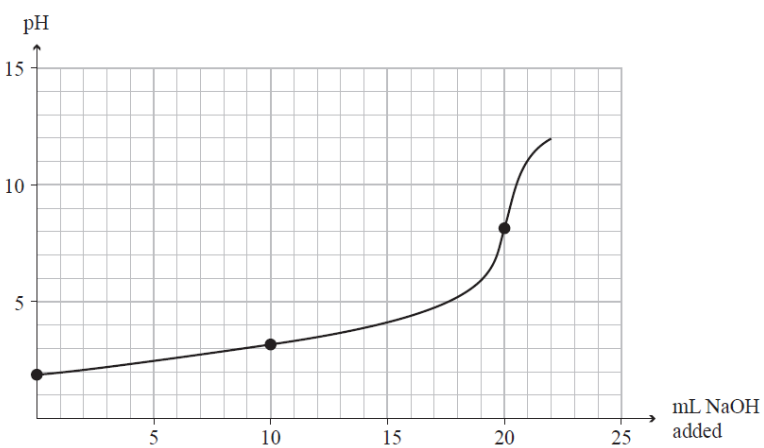
NCEA 2015 Titration Curves - (PART FOUR)

Excellence Question

Question: 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

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



$$n(\text{NaOH}) = cv = 0.258 \times \frac{24 - 20}{1000} = 1.032 \times 10^{-3} \text{ mol}$$

$$c(\text{NaOH}) = \frac{n}{v} = \frac{1.032 \times 10^{-3}}{44/1000} = 0.0235 \text{ mol L}^{-1}$$

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1 \times 10^{-14}}{0.0235} = 4.26 \times 10^{-13} \text{ mol L}^{-1}$$

$$\text{pH} = -\log 4.26 \times 10^{-13} = 12.4$$

NCEA 2015 Titration Curves - (PART FIVE)

Question: 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.

$pK_a(CH_3COOH) = 4.76$ No calculations are necessary.

NCEA 2015 Titration Curves - (PART FIVE)

Excellence Question

Question: 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.

→ larger pKa more reactants

Since CH_3COOH has a higher pK_a, it is a weaker acid than HF. Therefore its conjugate base, CH_3COO^-, will be a stronger base than F^-. This means $[OH^-]$ will be higher at the equivalence point for the CH_3COOH vs NaOH titration, so the equivalence point pH will be higher.

NCEA 2016 Titration Curves - (PART ONE)

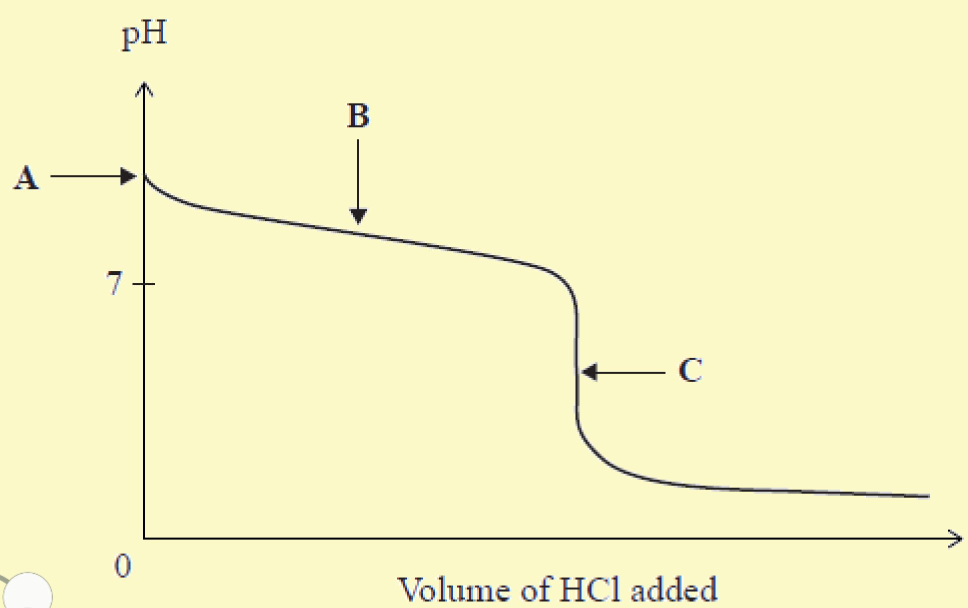
Question: 3a: 20.00 mL of 0.320 mol L^{-1} ammonia, NH_3, is titrated with 0.640 mol L^{-} hydrochloric acid, HCl.

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.



NCEA 2016 Titration Curves - (PART ONE)

Merit Question

Question: 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:

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

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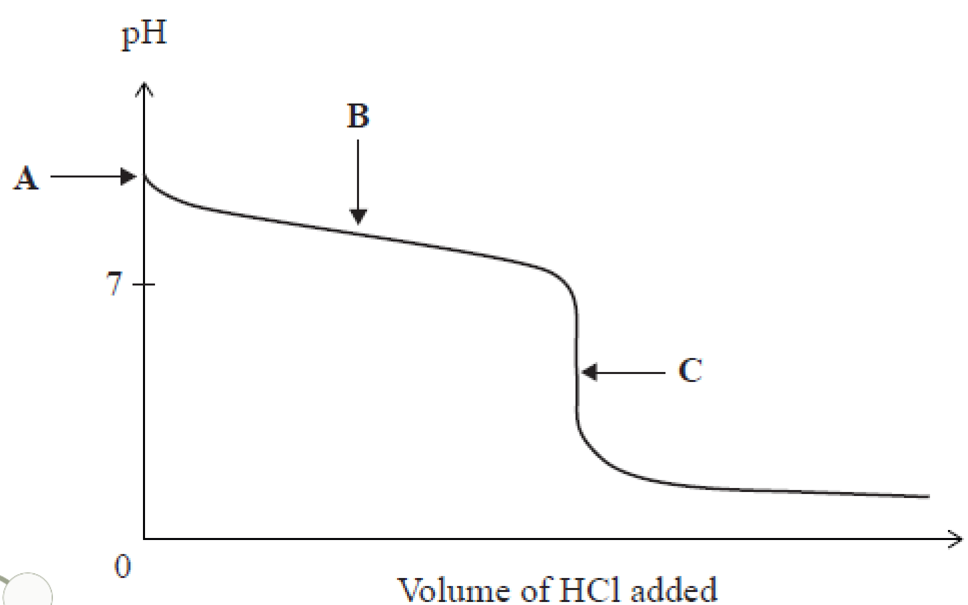


(Ammonium chloride) is acidic

$NH_4^+ + H_2O \leftrightharpoons H_3O^+ + NH_3$

So therefore

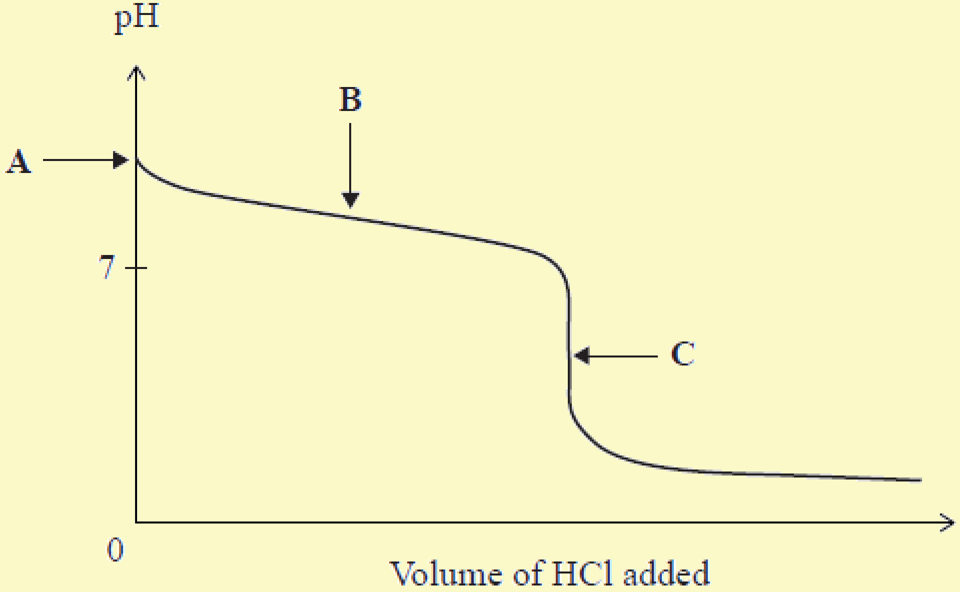
$[H_3O^+] > [OH^-]$

When considering pH think about which ion, H_3O^+ or OH^-, will be at the higher concentration

NCEA 2016 Titration Curves - (PART TWO)

Question 3c: 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.



NCEA 2016 Titration Curves - (PART TWO)

Excellence Question

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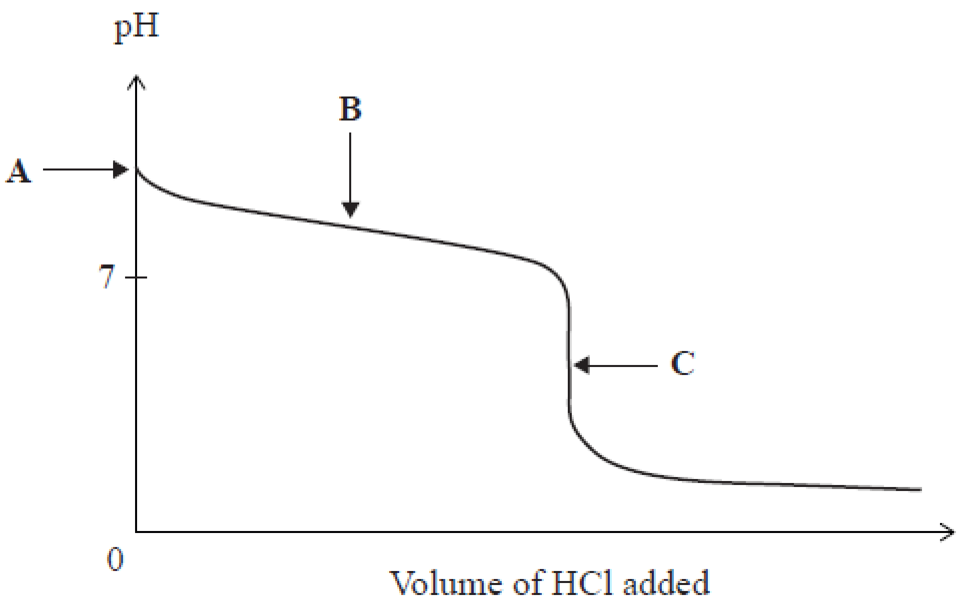


$[NH_4^+] = 0.320 \times 20 / 30 = 0.213\ molL^{-1}$

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

$[H_3O^+] = \sqrt{(5.75 \times 10^{-10} \times 0.213)}$

$= 1.11 \times 10^{-5}\ molL^{-1}$

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

$pH = 4.96$

NCEA 2016 Titration Curves - (PART THREE)

Question 3b: 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

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

NCEA 2016 Titration Curves - (PART THREE)

Excellence Question

Question 3b: 20.00 mL of 0.320 mol L^{-1} ammonia, NH_3, is titrated with 0.640 mol L^{-1} hydrochloric acid, HCl. $pK_a(NH_4^+) = 9.24$

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

Since B is half way to the equivalence point, $[NH_4^+] = [NH_3]$.

OR

$pK_a = pH + \log [acid] \div [c.base]$

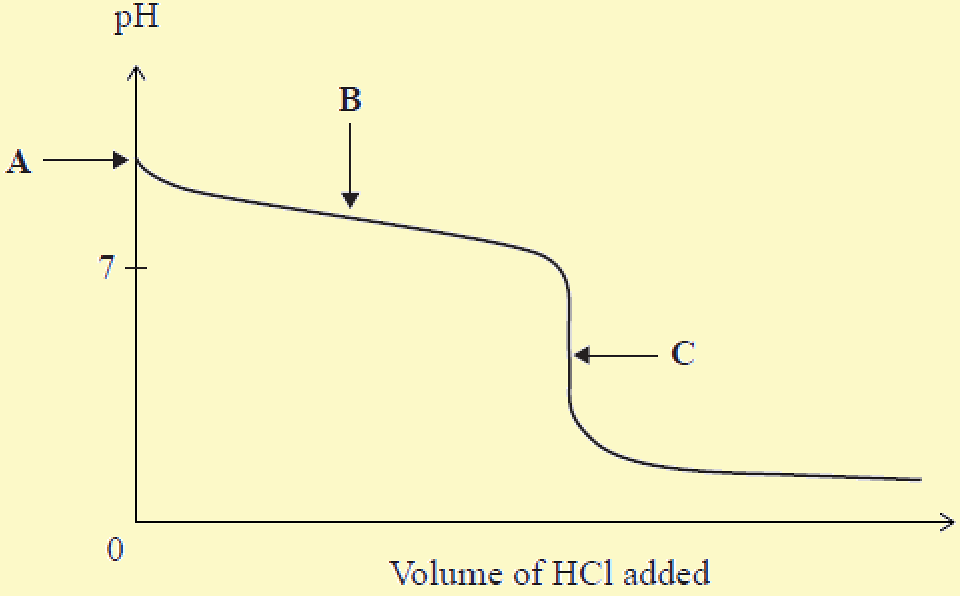
so $K_a = [H_3O^+]$

therefore $pK_a = pH$.

NCEA 2016 Titration Curves - (PART FOUR)

Question 3d: 20.00 mL of 0.320 mol L^{-1} ammonia, NH_3, is titrated with 0.640 mol L^{-1} hydrochloric acid, HCl. $pK_a(NH_4^+) = 9.24$

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

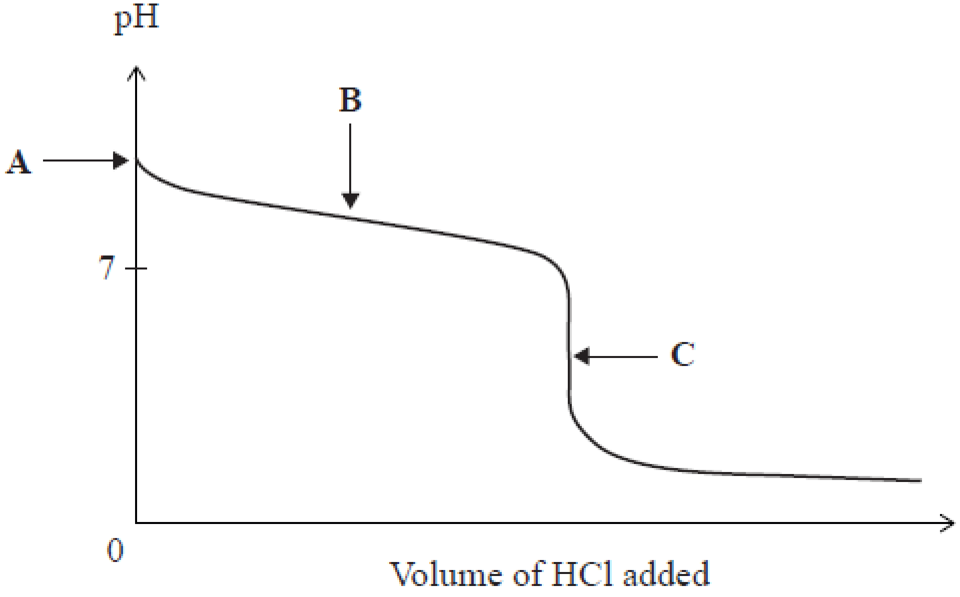


NCEA 2016 Titration Curves - (PART FOUR)

Merit Question

Question 3d: 20.00 mL of 0.320 mol L^{-1} ammonia, NH_3, is titrated with 0.640 mol L^{-1} hydrochloric acid, HCl. $pK_a(NH_4^+) = 9.24$

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



The solution at the equivalence point is NH_4Cl.

NH_4^+ solution is acidic since,

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$$

NCEA 2013 Buffers

Question: 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

NCEA 2013 Buffers

Excellence Question

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

$$K_a = \frac{[CH_3NH_2][H_3O^+]}{[CH_3NH_3^+]}$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} \quad \text{to} \quad [H_3O^+] = \frac{K_a \times [HA]}{[A^-]}$$

$$[H_3O^+] = \frac{K_a[CH_3NH_3^+]}{[CH_3NH_2]}$$

20mL in a 50mL total

30mL in a 50mL total

$$[CH_3NH_2] = \frac{30 \times 10^{-3} \times 1}{50 \times 10^{-3}} = 0.600 \text{ mol L}^{-1}$$

$$[CH_3NH_3^+] = \frac{20 \times 10^{-3} \times 1}{50 \times 10^{-3}} = 0.400 \text{ mol L}^{-1}$$

$$[H_3O^+] = 1.52705 \times 10^{-11} \text{ mol L}^{-1}$$

$$pH = 10.8$$

GZ Science Resources

NCEA 2013 Buffers

Question: 1c: 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

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

NCEA 2013 Buffers

Excellence Question

Question: 1c: 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

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

When a small amount of acid (H_3O^+) ions are added, they will react with the $CH_3NH_{2(aq)}$ molecules to form $CH_3NH_3{}^+{}_{(aq)}$ ions.

$$CH_3NH_{2(aq)} + H_3O^+{}_{(aq)} \rightarrow CH_3NH_3{}^+{}_{(aq)} + H_2O_{(\ell)}$$

The added acid (H_3O^+), is mostly consumed, and the pH of the solution changes very little.

NCEA 2014 Buffers

Question: 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

NCEA 2014 Buffers

Excellence Question

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Assume there is no change in volume.

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

$$K_a = \frac{[F^-][H_3O^+]}{[HF]}$$

$$10^{-3.17} = \frac{[F^-] \times 10^{-4.02}}{0.0500}$$

$$[F^-] = 0.354 \text{ mol L}^{-1}$$

$$n(NaF) = 0.354 \text{ mol L}^{-1} \times 0.150 \text{ L} = 0.0531 \text{ mol}$$

$$m(NaF) = 0.0531 \text{ mol} \times 42.0 \text{ g mol}^{-1} = 2.23 \text{ g}$$

n = c x *V*

m= n x M