Chemistry AS 91392
C3.6 Aqueous Systems
NCEA Revision 2013 - 2016
Question: 2a: In an experiment, a saturated solution was made by dissolving $1.44 \times 10^{-3}$ g of Ag$_2$CrO$_4$ in water, and making it up to a volume of 50.0 mL.

$M (\text{Ag}_2\text{CrO}_4) = 332$ g mol$^{-1}$

(a) Write the $K_s$ expression for Ag$_2$CrO$_4(s)$.
Question: 2a: In an experiment, a saturated solution was made by dissolving $1.44 \times 10^{-3}$ g of Ag$_2$CrO$_4$ in water, and making it up to a volume of 50.0 mL.

$M (\text{Ag}_2\text{CrO}_4) = 332$ g mol$^{-1}$

(a) Write the $K_s$ expression for Ag$_2$CrO$_4(s)$.

$K_s = [\text{Ag}^+]^2[\text{CrO}_4^{2-}]$
Question: 2b: In an experiment, a saturated solution was made by dissolving $1.44 \times 10^{-3}$ g of $\text{Ag}_2\text{CrO}_4$ in water, and making it up to a volume of 50.0 mL. $M (\text{Ag}_2\text{CrO}_4) = 332$ g mol$^{-1}$

(i) Calculate the solubility of $\text{Ag}_2\text{CrO}_4(s)$, and hence give the $[\text{Ag}^+]$ and $[\text{CrO}_4^{2-}]$ in the solution.

(ii) Determine the $K_s(\text{Ag}_2\text{CrO}_4)$. 
Question: 2b: In an experiment, a saturated solution was made by dissolving $1.44 \times 10^{-3}$ g of Ag₂CrO₄ in water, and making it up to a volume of 50.0 mL. $M(\text{Ag}_2\text{CrO}_4) = 332 \text{ g mol}^{-1}$

(i) Calculate the solubility of Ag₂CrO₄(s), and hence give the [Ag⁺] and [CrO₄²⁻] in the solution.

\[
K_s = [\text{Ag}^+]^2[\text{CrO}_4^{2-}]
\]

\[
n(\text{Ag}_2\text{CrO}_4) = \frac{1.44 \times 10^3}{332} = 4.33 \times 10^6 \text{ mol in 50 mL}
\]

\[
[\text{Ag}_2\text{CrO}_4] = \frac{4.33 \times 10^6}{50 \times 10^3} = 8.67 \times 10^5 \text{ mol L}^{-1}
\]

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

\[
[\text{CrO}_4^{2-}] = 8.67 \times 10^5 \text{ mol L}^{-1}
\]

(ii) Determine the $K_s(\text{Ag}_2\text{CrO}_4)$.

\[
K_s = (1.73 \times 10^4)^2 (8.67 \times 10^5)
\]

\[
= 2.61 \times 10^{12}
\]
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$_2$).
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$$\text{PbCl}_2(s) \rightleftharpoons \text{Pb}^{2+}(aq) + 2\text{Cl}^-(aq)$$

Question: 2a: (ii) Write the expression for $K_s$ (PbCl$_2$).

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

\[ K_s(\text{PbCl}_2) = 1.70 \times 10^{-5} \text{ at 25°C} \]
Question: 2a: (iii) Calculate the solubility (in mol L⁻¹) of lead(II) chloride in water at 25°C, and give the [Pb²⁺] and [Cl⁻] in the solution.

\(K_s(PbCl_2) = 1.70 \times 10^{-5}\) at 25°C

\[\text{[Pb}^2+\text{]} = x \quad \text{[Cl}^-\text{]} = 2x\]

\[K_s = 4x^3\]

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

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

\[= 1.62 \times 10^2\text{ mol L}^{-1}\]

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

\[\text{[Cl}^-\text{]} = 3.24 \times 10^{-1}\text{ mol L}^{-1}\]
Question: 2a: (i) Sufficient calcium carbonate, $\text{CaCO}_3(\text{s})$, is dissolved in water to make a saturated solution.

Write the equation for the equilibrium occurring in a saturated solution of $\text{CaCO}_3$.

Question: 2a: (ii) Write the expression for $K_s(\text{CaCO}_3)$. 
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$_3$.

\[
\text{CaCO}_3(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq)
\]

Question: 2a: (ii) Write the expression for $K_s$(CaCO$_3$).

\[
K_s = [\text{Ca}^{2+}][\text{CO}_3^{2-}]
\]
Question: 2a: (iii) Calculate the solubility product of CaCO$_3$, $K_s$(CaCO$_3$).

The solubility of CaCO$_3$ is $5.74 \times 10^{-5}$ mol L$^{-1}$. 
Question: 2a: (iii) Calculate the solubility product of CaCO$_3$, $K_s$(CaCO$_3$).

The solubility of CaCO$_3$ is $5.74 \times 10^{-5}$ mol L$^{-1}$.

$$K_s$(CaCO$_3$) = (5.74 \times 10^{-5})^2 = 3.29 \times 10^{-9}$$
**Question: 1a:** Silver carbonate, $\text{Ag}_2\text{CO}_3$, is a sparingly soluble salt.

$K_s(\text{Ag}_2\text{CO}_3) = 8.10 \times 10^{-12}$ at $25^\circ\text{C}$ \hspace{1cm} $M(\text{Ag}_2\text{CO}_3) = 276 \text{ g mol}^{-1}$

(a) Write the solubility product expression, $K_s$, for silver carbonate ($\text{Ag}_2\text{CO}_3$).
Question: 1a: Silver carbonate, Ag₂CO₃, is a sparingly soluble salt.

\[ K_s(\text{Ag}_2\text{CO}_3) = 8.10 \times 10^{-12} \text{ at 25ºC} \quad M(\text{Ag}_2\text{CO}_3) = 276 \text{ g mol}^{-1} \]

(a) Write the solubility product expression, \( K_s \), for silver carbonate (Ag₂CO₃).

\[ K_s = [\text{Ag}^+]^2[\text{CO}_3^{2-}] \]
Question: 1b: Silver carbonate, Ag$_2$CO$_3$, is a sparingly soluble salt.

$K_s(\text{Ag}_2\text{CO}_3) = 8.10 \times 10^{-12}$ at 25ºC \hspace{1cm} M(\text{Ag}_2\text{CO}_3) = 276 \text{ g mol}^{-1}$

Calculate the mass of Ag$_2$CO$_3$ that will dissolve in 50 mL of water to make a saturated solution at 25ºC.
**Question: 1b:** Silver carbonate, \( \text{Ag}_2\text{CO}_3 \), is a sparingly soluble salt.

\[ K_s(\text{Ag}_2\text{CO}_3) = 8.10 \times 10^{-12} \text{ at 25ºC} \quad M(\text{Ag}_2\text{CO}_3) = 276 \text{ g mol}^{-1} \]

Calculate the mass of \( \text{Ag}_2\text{CO}_3 \) that will dissolve in 50 mL of water to make a saturated solution at 25ºC.

Let \( s = \) solubility

\[
\begin{align*}
[\text{Ag}^+] &= 2s \\
[\text{CO}_3^{2-}] &= s \\
K_s &= 4s^3 \\
s &= 1.27 \times 10^{-4} \text{ mol L}^{-1} \\
n &= c \times v = 6.33 \times 10^{-6} \text{ mol} \\
m &= n \times M = 1.75 \times 10^{-3} \text{ g}
\end{align*}
\]

OR

\[
\begin{align*}
g \text{ L}^{-1} &= c \times M = 0.0349 \text{ g L}^{-1} \\
\text{so mass in 50 mL} &= \frac{0.0349 \times 50}{1000} \\
&= 1.75 \times 10^{-3} \text{ g}
\end{align*}
\]
Question: 2c: In another experiment, 0.0100 g of Ag₂CrO₄ in beaker A was made up to a volume of 50.0 mL with water. In beaker B, 0.0100 g of Ag₂CrO₄ was made up to a volume of 50.0 mL with 0.100 mol L⁻¹ ammonia solution.
Question: 2c: In another experiment, 0.0100 g of Ag\(_2\)CrO\(_4\) in beaker A was made up to a volume of 50.0 mL with water. In beaker B, 0.0100 g of Ag\(_2\)CrO\(_4\) was made up to a volume of 50.0 mL with 0.100 mol L\(^{-1}\) ammonia solution.

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 \(1.44 \times 10^{-3}\) 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\(_2\)CrO\(_4\) will dissociate completely and form an equilibrium.

\[
\text{Ag}_2\text{CrO}_4 \rightleftharpoons 2\text{Ag}^+ + \text{CrO}_4^{2-}
\]

\[
\text{Ag}^+ + 2\text{NH}_3 \rightarrow [\text{Ag(NH}_3)_2]^+
\]

The silver ion will then react further with NH\(_3\), removing it from the above equilibrium. Thus, more Ag\(_2\)CrO\(_4\) will dissolve to re-establish equilibrium.
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.
**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$_3$O$^+$ to form water,

H$_3$O$^+$ + OH$^-$ → H$_2$O  
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$^-$ → [Zn(OH)$_4$]$^{2-}$  OR Zn$^{2+}$ + 4OH$^-$ → [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.
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.
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$_3$O$^+$ 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.

$$2\text{H}_3\text{O}^+ + \text{CO}_3^{2-} \rightarrow 3\text{H}_2\text{O} + \text{CO}_2$$
Question: 1c: Explain how the solubility of Ag$_2$CO$_3$ will change if added to 50 mL of a 1.00 mol L$^{-1}$ ammonia, NH$_3$, solution.

Support your answer with balanced equations.
**Question: 1c:** Explain how the solubility of Ag\(_2\)CO\(_3\) will change if added to 50 mL of a 1.00 mol L\(^{-1}\) ammonia, NH\(_3\), solution.

Support your answer with balanced equations.

\[
\text{Ag}_2\text{CO}_3(s) \rightarrow 2\text{Ag}^+_{(aq)} + \text{CO}_3^{2-}_{(aq)}
\]

\[
\text{Ag}^+_{(aq)} + 2\text{NH}_3_{(aq)} \rightleftharpoons [\text{Ag(NH}_3)_2]_{(aq)}
\]

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**
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.

\[ K_s(\text{PbCl}_2) = 1.70 \times 10^{-5} \]
\[ M(\text{Pb(NO}_3)_2) = 331 \text{ g mol}^{-1} \]
**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.

\[K_s(\text{PbCl}_2) = 1.70 \times 10^{-5} \quad M(\text{Pb(NO}_3)_2) = 331 \text{ g mol}^{-1}\]

\[
n(\text{Pb(NO}_3)_2) = \frac{2.00 \text{ g}}{331 \text{ g mol}^{-1}}
\]

\[= 6.04 \times 10^{-3} \text{ mol}\]

\[\therefore [\text{Pb}^{2+}] = 6.04 \times 10^{-3} \text{ mol} / 0.500\text{L} = 1.21 \times 10^{-2} \text{ mol L}^{-1}\]

\[Q = (1.21 \times 10^{-2}) \times (0.440)^2 = 2.34 \times 10^{-3}\]

As \(Q > K_s\), a precipitate will form.
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.
**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

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

$Q$ = $[\text{Pb}^{2+}][\text{OH}^-]^2$

$[\text{Pb}^{2+}] = 0.5 \times 0.00421 = 2.105 \times 10^{-3}$

$[\text{OH}^-] = 0.5 \times 0.0398 = 1.99 \times 10^{-2}$

$Q = (2.105 \times 10^{-3}) \times (1.99 \times 10^{-2})^2$

$Q = 8.34 \times 10^{-7}$

Since $Q > K_s$, a precipitate of Pb(OH)$_2$ will form.
Question: 1d: Show by calculation whether a precipitate of Ag₂CO₃ will form when 20.0 mL of 0.105 mol L⁻¹ silver nitrate, AgNO₃, solution is added to 35.0 mL of a 0.221 mol L⁻¹ sodium carbonate, Na₂CO₃, solution.

\[ K_s(\text{Ag}_2\text{CO}_3) = 8.10 \times 10^{-12} \text{ at 25ºC} \]
**Question: 1d:** Show by calculation whether a precipitate of Ag$_2$CO$_3$ 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$_2$CO$_3$, solution.

$K_s$(Ag$_2$CO$_3$) = $8.10 \times 10^{-12}$ at 25ºC

\[
\text{AgNO}_3 \text{ dilution: } \frac{20}{55} \times 0.105 = 0.0382
\]

\[
\text{Na}_2\text{CO}_3 \text{ dilution: } \frac{35}{55} \times 0.221 = 0.141
\]

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

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

\[
K_s = [\text{Ag}^+]^2[\text{CO}_3^{2-}]
\]
2015 dissociation equations - NCEA Case Study

**Question: 1a: (i)** Methylammonium chloride, CH$_3$NH$_3$Cl, dissolves in water to form a weakly acidic solution.

$K_a(\text{CH}_3\text{NH}_3^+) = 2.29 \times 10^{-11}$

(a) (i) Write an equation to show CH$_3$NH$_3$Cl dissolving in water.

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2014 dissociation equations - NCEA Case Study

**Question: 1a:** When chlorine gas is added to water, the equation for the reaction is:

$$\text{Cl}_2(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HCl}_{(aq)} + \text{HOCl}_{(aq)}$$

- (i) Write an equation for the reaction of the weak acid, hypochlorous acid, HOCl, with water.
2015 dissociation equations - NCEA Case Study

Question: 1a: (i) Methylammonium chloride, \( \text{CH}_3\text{NH}_3\text{Cl} \), dissolves in water to form a weakly acidic solution.

\[ \text{Ka(\text{CH}_3\text{NH}_3^+}) = 2.29 \times 10^{-11} \]

(a) (i) Write an equation to show \( \text{CH}_3\text{NH}_3\text{Cl} \) dissolving in water.

\[
\text{CH}_3\text{NH}_3\text{Cl} (s) \rightarrow \text{CH}_3\text{NH}_3^+ (aq) + \text{Cl}^- (aq) \quad \text{CH}_3\text{NH}_3^+ + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_2 + \text{H}_3\text{O}^+ 
\]

2014 dissociation equations - NCEA Case Study

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

\[
\text{Cl}_2 (g) + \text{H}_2\text{O} (l) \rightleftharpoons \text{HCl} (aq) + \text{HOCl} (aq)
\]

(i) Write an equation for the reaction of the weak acid, hypochlorous acid, \( \text{HOCl} \), with water.

\[
\text{HOCl} + \text{H}_2\text{O} \rightleftharpoons \text{OCl}^- + \text{H}_3\text{O}^+
\]
Ethanamine, CH₃CH₂NH₂, is a weak base.

\[ pK_a(CH₃CH₂NH₃^+) = 10.6 \quad K_a(CH₃CH₂NH₃^+) = 2.51 \times 10^{-11} \]

(a) Write an equation to show the reaction of ethanamine with water.
Question: 2a: (i) Ethanamine, CH₃CH₂NH₂, is a weak base.

\[ pK_a(\text{CH}_3\text{CH}_2\text{NH}_3^+) = 10.6 \quad K_a(\text{CH}_3\text{CH}_2\text{NH}_3^+) = 2.51 \times 10^{-11} \]

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

\[ \text{CH}_3\text{CH}_2\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{CH}_2\text{NH}_3^+ + \text{OH}^- \]
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$_3$NH$_3$Cl  CH$_3$NH$_2$  HCl
**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.

CH₃NH₃Cl  CH₃NH₂  HCl

**HCl < CH₃NH₃Cl < CH₃NH₂**

- HCl, a strong acid, reacts completely with water to form 1 mol L⁻¹ H₃O⁺ and hence a low pH. 
  \[ \text{HCl} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^- \]

- CH₃NH₃Cl dissociates completely in water to form CH₃NH₃⁺ and Cl⁻. CH₃NH₃⁺, a weak acid, partially reacts with water to form less than 1 mol L⁻¹ H₃O⁺ and hence a higher pH than HCl. 
  \[ \text{CH}_3\text{NH}_3\text{Cl} \rightarrow \text{CH}_3\text{NH}_3^+ + \text{Cl}^- \]
  \[ \text{CH}_3\text{NH}_3^+ + \text{H}_2\text{O} \leftrightarrow \text{CH}_3\text{NH}_2 + \text{H}_3\text{O}^+ \]

- CH₃NH₂, a weak base, partially reacts with water to form OH⁻ ions. So there are more OH⁻ ions than H₃O⁺ ions and the pH is thus high. 
  \[ \text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \leftrightarrow \text{CH}_3\text{NH}_3^+ + \text{OH}^- \]
**Question: 1a:** When chlorine gas is added to water, the equation for the reaction is: \[ \text{Cl}_2(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HCl}(aq) + \text{HOCl}(aq) \]

(ii) List all the species present when HOCl reacts with water, in order of decreasing concentration. Justify your order.
**Question: 1a:** When chlorine gas is added to water, the equation for the reaction is: \[ \text{Cl}_2(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HCl}(aq) + \text{HOCl}(aq) \]

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

\[
\text{HOCl} > \text{H}_3\text{O}^+ > \text{OCl}^- > \text{OH}^- \quad \text{or} \quad \text{HOCl} > \text{H}_3\text{O}^+ = \text{OCl}^- > \text{OH}^-
\]

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

\(\text{H}_3\text{O}^+ \) and \(\text{OCl}^-\) are produced in equal amounts / there is a small contribution to \(\text{H}_3\text{O}^+\) from water therefore \(\text{H}_3\text{O}^+ > \text{OCl}^-\)

Because there is a relatively high \([\text{H}_3\text{O}^+]\), the \([\text{OH}^-]\) is very low (or links to \(K_w\)).
**Question: 2c:** Ethyl ammonium chloride, CH\textsubscript{3}CH\textsubscript{2}NH\textsubscript{3}Cl, is a weak acid that will also react with water.

List all the species present in a solution of CH\textsubscript{3}CH\textsubscript{2}NH\textsubscript{3}Cl, in order of decreasing concentration.

Do not include water.

Justify the order you have given.

Include equations, where necessary.
**Question: 2c:** Ethyl ammonium chloride, CH$_3$CH$_2$NH$_3$Cl, is a weak acid that will also react with water. List all the species present in a solution of CH$_3$CH$_2$NH$_3$Cl, in order of decreasing concentration. Do not include water. Justify the order you have given. Include equations, where necessary.

Cl$^-$ > CH$_3$NH$_3^+$ > H$_3$O$^+$ = CH$_3$NH$_2$ > OH$^-$  OR

Cl$^-$ > CH$_3$NH$_3^+$ > H$_3$O$^+$ > CH$_3$NH$_2$ > OH$^{-}$

CH$_3$CH$_2$NH$_3$Cl → CH$_3$CH$_2$NH$_3^+$ + Cl$^-$

CH$_3$CH$_2$NH$_3$Cl 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$_3$CH$_2$NH$_3^+$ + H$_2$O ⇌ CH$_3$CH$_2$NH$_2$ + H$_3$O$^+$

For every mole of CH$_3$CH$_2$NH$_3^+$ that reacts with water, 1 mole of CH$_3$CH$_2$NH$_2$ and H$_3$O$^+$ are formed.

(However, H$_3$O$^+$ is slightly more concentrated than CH$_3$CH$_2$NH$_2$, as there is a small contribution from water).

OH$^-$ is present in the lowest concentration as this comes from the dissociation of water only.
Question: 1b: The conductivity of the 1 mol L–1 solutions formed in (a) can be measured. CH$_3$NH$_3$Cl       CH$_3$NH$_2$       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.
**Question: 1b:** The conductivity of the 1 mol L⁻¹ solutions formed in (a) can be measured. CH₃NH₃Cl > CH₃NH₂ > 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.

\[ \text{HCl} = \text{CH₃NH₃Cl} > \text{CH₃NH₂} \]

CH₃NH₃Cl and HCl will dissociate completely in water to produce 2 mol L⁻¹ ions.

CH₃NH₂ will only partially react with water to produce less than 1 mol L⁻¹ of ions.
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.

<table>
<thead>
<tr>
<th>Solution</th>
<th>NaOH</th>
<th>CH₃NH₂</th>
<th>CH₃COONa</th>
</tr>
</thead>
<tbody>
<tr>
<td>pH</td>
<td>13.2</td>
<td>11.9</td>
<td>8.98</td>
</tr>
<tr>
<td>Electrical conductivity</td>
<td>good</td>
<td>poor</td>
<td>good</td>
</tr>
</tbody>
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<tbody>
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<td><strong>pH</strong></td>
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<tr>
<td><strong>Electrical conductivity</strong></td>
<td>good</td>
<td>poor</td>
<td>good</td>
</tr>
</tbody>
</table>

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 → Na⁺ + OH⁻
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.

<table>
<thead>
<tr>
<th>Solution</th>
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<th>CH₃NH₂</th>
<th>CH₃COONa</th>
</tr>
</thead>
<tbody>
<tr>
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</tbody>
</table>

- CH₃NH₂ is a weak base that partially reacts / dissociates / ionises with H₂O producing a lower concentration of OH⁻, Therefore it has a lower pH than NaOH:

CH₃NH₂ + H₂O ⇌ CH₃NH₃⁺ + OH⁻

- The CH₃COONa is an ionic solid that dissociates completely in H₂O. The CH₃COO⁻ ion is a weak base that partially reacts / dissociates / ionises with H₂O producing a lower concentration of OH⁻.

CH₃COO⁻ + H₂O ⇌ CH₃COOH + OH⁻

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

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**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 → Na⁺ + OH⁻  Therefore it is a good conductor.
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.

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Electrical conductivity:

- Since CH₃NH₂ 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.

  \[ \text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_3^+ + \text{OH}^- \]

- CH₃COONa is also an ionic solid. It dissolves completely to produce a high concentration of Na⁺ and CH₃COO⁻ ions:

  \[ \text{CH}_3\text{COONa} \rightarrow \text{Na}^+ + \text{CH}_3\text{COO}^- \]

Therefore it is a good conductor.
Question: 1a: Hypochlorous acid has a pK\textsubscript{a} of 7.53. Another weak acid, hydrofluoric acid, HF, has a pK\textsubscript{a} of 3.17.

A 0.100 mol L\textsuperscript{–1} solution of each acid was prepared by dissolving it in water. Compare the pHs of these two solutions.

*No calculations are necessary.*
Question: 1a: Hypochlorous acid has a $pK_a$ of 7.53. Another weak acid, hydrofluoric acid, HF, has a $pK_a$ 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 $pK_a$ more reactants, the weaker the acid
Question: 1a: (iv) Calculate the pH of 0.0152 mol L$^{-1}$ CH$_3$NH$_3$Cl solution.

$K_a(CH_3NH_3^+) = 2.29 \times 10^{-11}$
Question: 1a: (iv) Calculate the pH of 0.0152 mol L\(^{-1}\) CH\(_3\)NH\(_3\)Cl solution.

\(K_a(\text{CH}_3\text{NH}_3^+) = 2.29 \times 10^{-11}\)

\[
\left[\text{H}_3\text{O}^+\right] = \sqrt{K_a} \times \text{HA} \\
= 5.90 \times 10^{-7}
\]

\[\text{pH} = 6.23\]

\([\text{H}_3\text{O}^+] = 5.90 \times 10^{-7} \text{ mol L}^{-1}\]

\[\text{pH} = -\log 5.90 \times 10^{-7} = 6.23\]
Question: 2b: Calculate the pH of a 0.109 mol L⁻¹ solution of ethanamine.

\[ pK_a(\text{CH}_3\text{CH}_2\text{NH}_3^+) = 10.6 \]
Question: 2b: Calculate the pH of a 0.109 mol L\(^{-1}\) solution of ethanamine.

\[ pK_a (\text{CH}_3\text{CH}_2\text{NH}_3^+) = 10.6 \]

\[
[H_3O^+] = \sqrt{K_a \times K_w \div [\text{CH}_3\text{CH}_2\text{NH}_2]} \\
[H_3O^+] = \sqrt{2.51 \times 10^{-11} \times 1.00 \times 10^{-14} \div 0.109} \\
[H_3O^+] = 1.52 \times 10^{-12} \text{ mol L}^{-1}
\]

\[
\text{pH} = -\log [H_3O^+] = 11.8
\]

\[
K_b = \frac{1 \times 10^{-14}}{K_a}
\]

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

\[
[H_3O^+] = 1 \times 10^{-14} / [\text{OH}^-]
\]

\[
\text{pH} = -\log[H_3O^+]
\]
Question: 3a: 20.0 mL of 0.0896 mol L\(^{-1}\) ethanoic acid is titrated with 0.100 mol L\(^{-1}\) sodium hydroxide. \(pK_a\) (CH\(_3\)COOH) = 4.76

(a) Calculate the pH of the ethanoic acid before any NaOH is added.
Question: 3a: 20.0 mL of 0.0896 mol L\(^{-1}\) ethanoic acid is titrated with 0.100 mol L\(^{-1}\) sodium hydroxide. \(pK_a\) (CH\(_3\)COOH) = 4.76

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

\[
K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} \quad \text{• } \quad pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]}
\]

\[
[H_3O^+] = \sqrt{1.74 \times 10^{-5} \times 0.0896 \text{ mol L}^{-1}} \quad = 1.25 \times 10^{-3} \text{ mol L}^{-1}
\]

\[
pH = \log[H_3O^+] = 2.90
\]
**Question: 3b:** Halfway to the equivalence point of the titration, the pH = pKₐ of the ethanoic acid.

Discuss the reason for this.
**Question: 3b:** Halfway to the equivalence point of the titration, the pH = pKa 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 \([\text{CH}_3\text{COOH}] = [\text{CH}_3\text{COO}^-]\) then \(K_a = [\text{H}_3\text{O}^+]\)

So \(pK_a = pH\)
**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.
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Your answer should include chemical equations.

No calculations are required.

\[
\text{NaOH}(aq) + \text{CH}_3\text{COOH}(aq) \rightarrow \text{NaCH}_3\text{COO}(aq) + \text{H}_2\text{O}(l)
\]

- [CH$_3$COO$^-$] increases as it is formed in reaction
- [Na$^+$] increases as NaOH is added
- [CH$_3$COOH] decreases as it reacts with NaOH
- [H$_3$O$^+$] decreases because [CH$_3$COO$^-$] / [CH$_3$COOH] increases and $K_a$ is a constant.
- [OH$^-$] increases because [H$_3$O$^+$] decreases and [H$_3$O$^+$][OH$^-$] is constant.
Question: 3c: 20.0 mL of 0.0896 mol L\(^{-1}\) ethanoic acid is titrated with 0.100 mol L\(^{-1}\) sodium hydroxide. \(pK_a (\text{CH}_3\text{COOH}) = 4.76\)

(ii) Calculate the pH of the titration mixture after 5.00 mL of NaOH has been added.
Question: 3c: 20.0 mL of 0.0896 mol L\(^{-1}\) ethanoic acid is titrated with 0.100 mol L\(^{-1}\) sodium hydroxide. \(pK_a\) (CH\(_3\)COOH) = 4.76

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

\[
\text{NaOH}(aq) + \text{CH}_3\text{COOH}_{(aq)} \rightarrow \text{NaCH}_3\text{COO}_{(aq)} + \text{H}_2\text{O}_{(l)}
\]

\[
n(\text{CH}_3\text{COOH at start}) = 0.0896 \times (20 \times 10^{-3}) = 1.79 \times 10^{-3} \text{ mol} \\
n(\text{NaOH added}) = 0.1 \times (5 \times 10^{-3}) = 5 \times 10^{-4} \text{ mol}
\]

After 5 mL NaOH added: (total 25mL)

\[
n(\text{CH}_3\text{COOH}) = 1.29 \times 10^{-3} \text{ mol}
\]

\[
n(\text{CH}_3\text{COO}^-) = 5 \times 10^{-4} \text{ mol}
\]

\[
[\text{CH}_3\text{COOH}] = 0.0516 \text{ mol L}^{-1} \\
[\text{CH}_3\text{COO}^-] = 0.0200 \text{ mol L}^{-1} \\
[H_3O^+] = 4.48 \times 10^{-5} \text{ mol L}^{-1}
\]

\[
\text{pH} = 4.35
\]

**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 mol of base required to react with acid produces 1 mol of conjugate base.

Example:

\[
\text{HCOOH}_{(aq)} + \text{NaOH}_{(aq)} \leftrightarrow \text{HCOONa}_{(aq)} + \text{H}_2\text{O}_{(l)}
\]

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

\[
V = \text{initial volume in flask} + \text{volume added during titration to reach equivalence}
\]

c) Use pH equations

\[
[H_3O^+] = \sqrt{\frac{k_s \times k_w}{c(\text{conjugate base})}}
\]
Question: 3a: A titration was carried out by adding hydrobromic acid, HBr, to 20.0 mL of aqueous methylamine, CH₃NH₂, solution.

The equation for the reaction is:

\[ \text{CH}_3\text{NH}_2 + \text{HBr} \rightarrow \text{CH}_3\text{NH}_3^+ + \text{Br}^- \]

\[ K_a(\text{CH}_3\text{NH}_3^+) = 2.29 \times 10^{-11} \]

Explain why the pH does not change significantly between the addition of 5 to 15 mL of HBr (around point A on the curve).
**Question: 3a:** A titration was carried out by adding hydrobromic acid, HBr, to 20.0 mL of aqueous methylamine, CH₃NH₂, solution.

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\[
K_a(\text{CH}_3\text{NH}_3^+) = 2.29 \times 10^{-11}
\]

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, \([\text{CH}_3\text{NH}_2] \approx [\text{CH}_3\text{NH}_3^+].\) So the solution has buffering properties in the proximity of point A. When HBr is added, the \(\text{H}_3\text{O}^+\) is consumed:

\[
\text{H}_3\text{O}^+ + \text{CH}_3\text{NH}_2 \rightarrow \text{CH}_3\text{NH}_3^+ + \text{H}_2\text{O}
\]

Since the \(\text{H}_3\text{O}^+\) is removed from the solution (neutralised), the pH does not change significantly.
Question: 3b: A titration was carried out by adding hydrobromic acid, HBr, to 20.0 mL of aqueous methylamine, CH₃NH₂, solution.

The equation for the reaction is: CH₃NH₂ + HBr → CH₃NH₃+ + Br⁻

\[ K_a(CH₃NH₃^+) = 2.29 \times 10^{-11} \]

The aqueous methylamine, CH₃NH₂, 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.
Question: 3b: A titration was carried out by adding hydrobromic acid, HBr, to 20.0 mL of aqueous methylamine, CH$_3$NH$_2$, solution.

The equation for the reaction is: CH$_3$NH$_2$ + HBr → CH$_3$NH$_3^+$ + Br$^-$

$K_a$(CH$_3$NH$_3^+$) = 2.29 × 10$^{-11}$

The aqueous methylamine, CH$_3$NH$_2$, 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$^{-1}$.

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

\[
K_a = \frac{[CH_3NH_2][H_3O^+]}{[CH_3NH_3^+]} = \frac{[OH^-]}{[CH_3NH_2][H_3O^+]} = \frac{2.29 \times 10^{-11}}{1 \times 10^{-14}}
\]

\[
[CH_3NH_2] = \frac{(2.29 \times 10^{-11}) \times (1 \times 10^{-14})}{(10^{-11.8})^2} = 0.0912 \text{ mol L}^{-1}
\]
**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.
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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.

\[ \text{CH}_3\text{NH}_3^+, \text{Br}^-, \text{CH}_3\text{NH}_2, \text{H}_3\text{O}^+ \]
Question: 3c: (ii) Compare and contrast the solution at point B with the initial aqueous methylamine solution.
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$_3$NH$_2$) which only partially reacts with water to produce a relatively low concentration of ions.

As a result, the initial CH$_3$NH$_2$ solution will be a poor electrical conductor.

$$\text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_3^+ + \text{OH}^-$$

Therefore species present are $\text{CH}_3\text{NH}_2 > \text{OH}^- \geq \text{CH}_3\text{NH}_3^+ > \text{H}_3\text{O}^+$

At point B, there is a solution of the salt CH$_3$NH$_3$Br present which is dissociated completely into ions. Therefore there is a relatively high concentration of ions (CH$_3$NH$_3^+$ and Br$^-$) present in the solution, so it will be a good electrical conductor / electrolyte.

$$\text{CH}_3\text{NH}_3\text{Br} \rightarrow \text{CH}_3\text{NH}_3^+ + \text{Br}^-$$

CH$_3$NH$_3^+$ reacts with water according to the equation

$$\text{CH}_3\text{NH}_3^+ + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_2 + \text{H}_3\text{O}^+$$

Species present are $\text{Br}^- > \text{CH}_3\text{NH}_3^+ > \text{H}_3\text{O}^+ \geq \text{CH}_3\text{NH}_2 > (\text{OH}^-)$
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:

\[
\text{HF} + \text{NaOH} \rightarrow \text{NaF} + \text{H}_2\text{O}
\]

\(pK_a(\text{HF}) = 3.17\)

(i) Identify the species in solution at the equivalence point.
Question: 3a: 20.0 mL of 0.258 mol L$^{-1}$ hydrofluoric acid, HF, solution is titrated with a sodium hydroxide, NaOH, solution.

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\[pK_a(\text{HF}) = 3.17\]

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

\[\text{Na}^+, \text{F}^-, \text{H}_2\text{O}, \text{HF}, \text{OH}^-, \text{H}_3\text{O}^+.\]
Question: 3a: (ii) Explain why the pH at the equivalence point is greater than 7. Include an equation in your answer.
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.
Question: 3a: 20.0 mL of 0.258 mol L\(^{-1}\) hydrofluoric acid, HF, solution is titrated with a sodium hydroxide, NaOH, solution.

\[
\text{HF} + \text{NaOH} \rightarrow \text{NaF} + \text{H}_2\text{O}
\]

\(pK_a(\text{HF}) = 3.17\)

(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.
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 → NaF + H₂O  \( pK_a(HF) = 3.17 \)

(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.

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

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

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

However, since the pH of the buffer solution is less than the pKa, \( [HF] > [F^-] \), it is more effective against added base than acid.
Question: 3a: 20.0 mL of 0.258 mol L\(^{-1}\) hydrofluoric acid, HF, solution is titrated with a sodium hydroxide, NaOH, solution.

\[
\text{HF} + \text{NaOH} \rightarrow \text{NaF} + \text{H}_2\text{O}
\]

\(p\text{K}_a(\text{HF}) = 3.17\)

(iv) Determine by calculation, the pH of the solution after 24.0 mL of 0.258 mol L\(^{-1}\) NaOH solution has been added.
**Question: 3a:** 20.0 mL of 0.258 mol L\(^{-1}\) hydrofluoric acid, HF, solution is titrated with a sodium hydroxide, NaOH, solution.

\[
\text{HF} + \text{NaOH} \rightarrow \text{NaF} + \text{H}_2\text{O}
\]

\(pK_a(\text{HF}) = 3.17\)

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

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

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

\[
[H_3O^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1 \times 10^{-14}}{0.0235} = 4.26 \times 10^{-13}\ \text{mol L}^{-1}
\]

\[
pH = \log 4.26 \times 10^{-13} = 12.4
\]
Question: 3b: In a second titration, a 0.258 mol L$^{-1}$ ethanoic acid, CH$_3$COOH, solution was titrated with the NaOH solution.

Contrast the expected pH at the equivalence point with the HF titration.

$pK_a$(CH$_3$COOH) = 4.76 No calculations are necessary.
Question: 3b: In a second titration, a 0.258 mol L\(^{-1}\) ethanoic acid, CH\(_3\)COOH, solution was titrated with the NaOH solution.

Contrast the expected pH at the equivalence point with the HF titration.

\[ pK_a(\text{CH}_3\text{COOH}) = 4.76 \quad \text{No calculations are necessary.} \]

Since CH\(_3\)COOH has a higher p\(K_a\), it is a weaker acid than HF. Therefore its conjugate base, CH\(_3\)COO\(^-\), will be a stronger base than F\(^-\). This means \([\text{OH}^-]\) will be higher at the equivalence point for the CH\(_3\)COOH vs NaOH titration, so the equivalence point pH will be higher.
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:

\[ \text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4^+ + \text{Cl}^- \quad pK_a(\text{NH}_4^+) = 9.24 \]

The curve for this titration is given below.

Explain why the pH at the equivalence point (point C) is not 7.
**Question: 3a:** 20.00 mL of 0.320 mol L\(^{-1}\) ammonia, \(\text{NH}_3\), is titrated with 0.640 mol L\(^{-1}\) hydrochloric acid, HCl.

The equation for this reaction is:

\[
\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4^+ + \text{Cl}^- \quad \text{p}K_a(\text{NH}_4^+) = 9.24
\]

The curve for this titration is given below.

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

(Ammonium chloride) is acidic

\[
\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NH}_3
\]

So therefore

\[
[\text{H}_3\text{O}^+] > [\text{OH}^-]
\]

When considering pH think about which ion, \(\text{H}_3\text{O}^+\) or \(\text{OH}^-\), will be at the higher concentration.
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.
Question 3c: 20.00 mL of 0.320 mol L\(^{-1}\) ammonia, \(\text{NH}_3\), is titrated with 0.640 mol L\(^{-1}\) hydrochloric acid, HCl. \(pK_a(\text{NH}_4^+) = 9.24\)

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

\[
[\text{NH}_4^+] = \frac{0.320 \times 20}{30} = 0.213 \text{ molL}^{-1}
\]

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

\[
[\text{H}_3\text{O}^+] = \sqrt{(5.75 \times 10^{-10} \times 0.213)}
\]

\[
= 1.11 \times 10^{-5} \text{ molL}^{-1}
\]

\[
\text{pH} = -\log[\text{H}_3\text{O}^+]
\]

\[
\text{pH} = 4.96
\]
**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.
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

\[pK_a = pH + \log \frac{[acid]}{[c.base]}\]

so \(K_a = [H_3O^+]\)

therefore \(pK_a = pH\).
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.
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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\(_4\)Cl.

NH\(_4^+\) solution is acidic since,

\[
\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+
\]
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\(_3\)NH\(_3\)Cl and 30.0 mL of 1 mol L\(^{-1}\) CH\(_3\)NH\(_2\)

Calculate the pH of the resultant buffer solution. \(pK_a (CH_3NH_3^+) = 10.64\)
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Calculate the pH of the resultant buffer solution. \(pK_a (CH_3NH_3^+) = 10.64\)

\[
K_a = \frac{[CH_3NH_2][H_3O^+]}{[CH_3NH_3^+]} \quad \text{to} \quad [H_3O^+] = K_a \times \frac{[HA]}{[A^-]}
\]

\[
[H_3O^+] = \frac{K_a [CH_3NH_3^+]}{[CH_3NH_2]} \quad [CH_3NH_2] = \frac{30 \times 10^3}{50 \times 10^3} = 0.600 \text{ mol L}^{-1}
\]

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

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

pH = 10.8
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20.0 mL of 1 mol L\(^{-1}\) CH\(_3\)NH\(_3\)Cl and 30.0 mL of 1 mol L\(^{-1}\) CH\(_3\)NH\(_2\)
(ii) Explain the effect on the solution formed in (i) when a small amount of acid is added.
**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\(_3\)NH\(_3\)Cl and 30.0 mL of 1 mol L\(^{-1}\) CH\(_3\)NH\(_2\)

(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\(_3\)O\(^+\)) ions are added, they will react with the CH\(_3\)NH\(_2(aq)\) molecules to form CH\(_3\)NH\(_3^+\)(aq) ions.

\[
\text{CH}_3\text{NH}_2(aq) + \text{H}_3\text{O}^+(aq) \rightarrow \text{CH}_3\text{NH}_3^+(aq) + \text{H}_2\text{O}(\ell)
\]

The added acid (H\(_3\)O\(^+\)), is mostly consumed, and the pH of the solution changes very little.
**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.

\[ M(\text{NaF}) = 42.0 \text{ g mol}^{-1} \quad pK_a(\text{HF}) = 3.17 \]
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\[ M(\text{NaF}) = 42.0 \text{ g mol}^{-1} \quad pK_a(\text{HF}) = 3.17 \]

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

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

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

\[
n(\text{NaF}) = 0.354 \text{ mol L}^{-1} \quad 0.150 \text{ L} = 0.0531 \text{ mol}
\]

\[
m(\text{NaF}) = 0.0531 \text{ mol} \quad 42.0 \text{ g mol}^{-1} = 2.23 \text{ g}
\]