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91392



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Mana Tohu Mātauranga o Aotearoa
New Zealand Qualifications Authority

Level 3 Chemistry 2025

91392 Demonstrate understanding of equilibrium principles in aqueous systems

Credits: Five

Achievement	Achievement with Merit	Achievement with Excellence
Demonstrate understanding of equilibrium principles in aqueous systems.	Demonstrate in-depth understanding of equilibrium principles in aqueous systems.	Demonstrate comprehensive understanding of equilibrium principles in aqueous systems.

Check that the National Student Number (NSN) on your admission slip is the same as the number at the top of this page.

You should attempt ALL the questions in this booklet.

A periodic table and other reference material are provided in the Resource Booklet L3-CHEMR.

If you need more room for any answer, use the extra space provided at the back of this booklet.

Check that this booklet has pages 2–12 in the correct order and that none of these pages is blank.

Do not write in the margins (⚡⚡⚡). This area will be cut off when the booklet is marked.

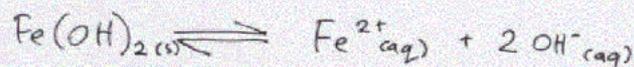
YOU MUST HAND THIS BOOKLET TO THE SUPERVISOR AT THE END OF THE EXAMINATION.

Merit

TOTAL 16

QUESTION ONE

- (a) (i) Write the equation for the equilibrium occurring in a saturated solution of iron(II) hydroxide, $\text{Fe}(\text{OH})_2$.



- (ii) Write the expression for $K_s(\text{Fe}(\text{OH})_2)$.

$$K_s(\text{Fe}(\text{OH})_2) = [\text{Fe}^{2+}][\text{OH}^{-}]^2$$

- (iii) Calculate the solubility product, K_s , of $\text{Fe}(\text{OH})_2$ in water at 25°C , given $\text{Fe}(\text{OH})_2$ has a solubility of $1.01 \times 10^{-5} \text{ mol L}^{-1}$.

let solubility = s

$$[\text{Fe}^{2+}] = s \quad [\text{OH}^{-}] = 2s$$

$$K_s = s \times 2s^2 \quad K_s = 4s^3$$

$$K_s = 4(1.01 \times 10^{-5})^3$$

$$1.01 \times 10^{-5} = 4s^3$$

$$s = \sqrt[3]{\frac{1.01 \times 10^{-5}}{4}}$$

$$s = 0.0136 \text{ mol L}^{-1} \quad (3 \text{ sf})$$

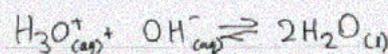
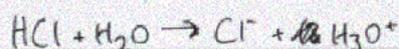
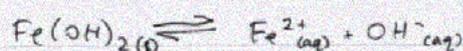
$$K_s = 4.12 \times 10^{-15}$$

- (iv) Some dilute hydrochloric acid, $\text{HCl}(\text{aq})$, is added to a saturated solution of $\text{Fe}(\text{OH})_2$.

Justify what will happen to the solubility of the $\text{Fe}(\text{OH})_2$ in solution.

Include relevant equation(s) in your answer.

No calculations are necessary.



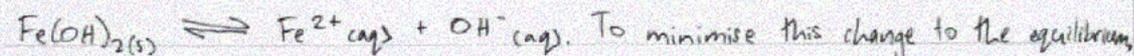
HCl is a strong acid and fully dissociates

into its ions, Cl^{-} and H_3O^{+} . When

added to a saturated solution of

$\text{Fe}(\text{OH})_2$, the H_3O^{+} will neutralise

the OH^{-} . This will lower the concentration of OH^{-} in the equilibrium:

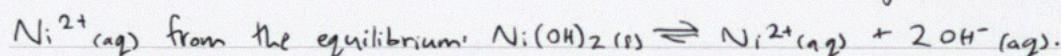
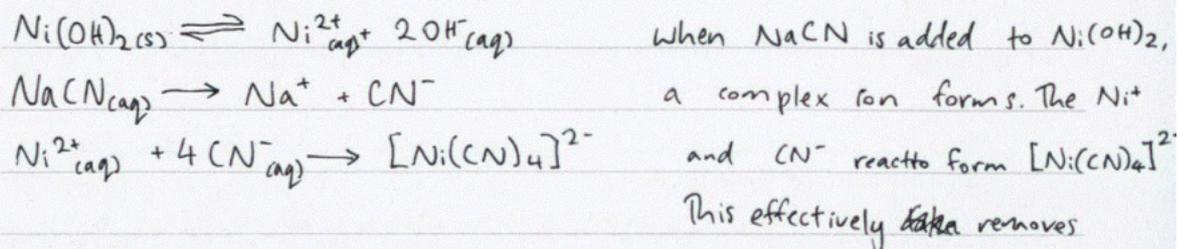


To minimise this change to the equilibrium, the forwards reaction will be favoured to replace some of the OH^{-} . Therefore more $\text{Fe}(\text{OH})_2(\text{s})$ will dissolve, increasing its solubility.

(b) Nickel hydroxide, $\text{Ni}(\text{OH})_2$, is another sparingly soluble solid.

- (i) A student makes a green precipitate of $\text{Ni}(\text{OH})_2$ in a test tube by mixing sodium hydroxide, $\text{NaOH}(\text{aq})$, and nickel chloride, $\text{NiCl}_2(\text{aq})$. When the student then adds excess sodium cyanide, $\text{NaCN}(\text{aq})$, the green precipitate dissolves.

Explain why the green precipitate of $\text{Ni}(\text{OH})_2$ dissolves, using equilibrium principles and any relevant equation(s).



To minimise this change to the equilibrium, the forwards reaction will be favoured. More $\text{Ni}(\text{OH})_2(\text{s})$ will dissolve into its ion to replace some of the $\text{Ni}^{2+}(\text{aq})$. Therefore the solubility of $\text{Ni}(\text{OH})_2$ increases ~~and~~ and the green precipitate of $\text{Ni}(\text{OH})_2$ dissolves.

- (ii) Determine whether a precipitate of $\text{Ni}(\text{OH})_2$ will form when 55.0 mL of 0.130 mol L^{-1} nickel sulfate, NiSO_4 , is added to 35.0 mL of potassium hydroxide, KOH , solution of pH 11.7.

$$K_s(\text{Ni}(\text{OH})_2) = 6 \times 10^{-16}$$

$$IP = [\text{Ni}^{2+}][\text{OH}^{-}]^2$$

$$[\text{Ni}^{2+}] = 0.130 \times \frac{55}{90} = 0.0794 \text{ mol L}^{-1}$$

$$[\text{OH}^{-}] = 2 \times$$

$$[\text{OH}^{-}] = 10^{-\text{pOH}} = 10^{-(14-11.7)} = 10^{-2.3} = 5.01 \times 10^{-3}$$

$$[\text{OH}^{-}] = 2 \times \therefore [\text{OH}^{-}] = 2 \times 5.01 \times 10^{-3} = 0.0100 \text{ mol L}^{-1}$$

$$IP = (0.0794) \times (0.0100)^2 = 7.98 \times 10^{-6}$$

$$K_s = 6 \times 10^{-16}$$

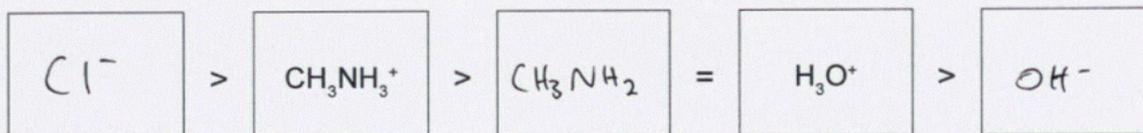
$$IP > K_s \quad 7.98 \times 10^{-6} > 6 \times 10^{-16} \quad \text{Since IP is greater than } K_s,$$

this means the solution is supersaturated and a precipitate will form.

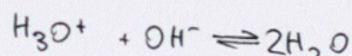
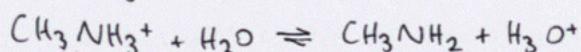
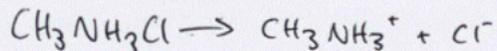
QUESTION TWO

Methylammonium chloride, $\text{CH}_3\text{NH}_3\text{Cl}$, is an acidic salt.

- (a) List all the species present in a solution of $\text{CH}_3\text{NH}_3\text{Cl}$ in order of decreasing concentration. Do not include water.



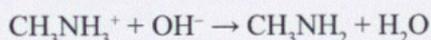
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- (b) If $\text{CH}_3\text{NH}_3\text{Cl}$ and methanamine, CH_3NH_2 , are mixed in the appropriate quantities, a buffer solution is formed.

$$K_a(\text{CH}_3\text{NH}_3^+) = 2.51 \times 10^{-11} \quad \text{p}K_a(\text{CH}_3\text{NH}_3^+) = 10.6$$

- (i) When a small volume of dilute sodium hydroxide, NaOH , is added to the buffer, the following reaction occurs:



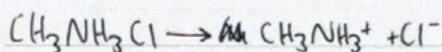
Describe the function of a buffer solution, and explain the significance of this equation in terms of the function of the buffer solution.

Buffer solutions are effective because they contain acidic and basic components so they can neutralise any added solution and resist changing pH.

The acidic component of the buffer: CH_3NH_3^+ is able to neutralise the OH^- from the added NaOH and resist a change to the pH.

- (ii) Calculate the mass of $\text{CH}_3\text{NH}_3\text{Cl}$ that should be added to 600 mL of $0.840 \text{ mol L}^{-1} \text{CH}_3\text{NH}_2$ to produce a solution of pH 12.1.

$$M(\text{CH}_3\text{NH}_3\text{Cl}) = 67.5 \text{ g mol}^{-1}$$



$$n = 0.0159$$



$$V = 600 \text{ mL} = 0.6 \text{ L}$$

$$\text{pH} = 12.1$$

$$C = 0.0266 \text{ mol L}^{-1}$$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-12.1} = 7.94 \times 10^{-13} \text{ mol L}^{-1}$$

$$M = 67.5 \text{ g mol}^{-1}$$

$$K_a = \frac{[\text{CH}_3\text{NH}_2][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{NH}_3^+]}$$

$$n = CV = 0.0266 \times 0.6 = 0.0159$$

$$[\text{CH}_3\text{NH}_3^+] = \frac{[\text{CH}_3\text{NH}_2][\text{H}_3\text{O}^+]}{K_a}$$

$$n = \frac{m}{M} \quad m = nM$$

$$[\text{CH}_3\text{NH}_3^+] = \frac{[\text{CH}_3\text{NH}_2][\text{H}_3\text{O}^+]}{K_a}$$

$$m = 0.0159 \times 67.5$$

$$m = 1.076 \text{ g}$$

$$[\text{CH}_3\text{NH}_3^+] = \frac{0.840 \times (7.94 \times 10^{-13})}{2.51 \times 10^{-11}}$$

$$m = 1.08 \text{ g (3sf)}$$

$$[\text{CH}_3\text{NH}_3^+] = 0.02657 = 0.0266 \text{ mol L}^{-1} \text{ (3sf)}$$

$$0.0266 \frac{\text{mol}}{1000 \text{ mL}}$$

$$0.6 \times 0.0266 = \frac{\text{mol}}{600 \text{ mL}} = 0.0159 \text{ mol / 600 mL}$$

- (iii) Explain why the solution prepared in part (ii) will not function as a buffer.

Outline how the solution could be modified in the laboratory to make it a buffer.

No calculations are necessary.

For a solution to function as a buffer, the pH must = $\text{p}K_a \pm 1$. $\text{p}K_a$ must be between 9.6 and 11.6. Because the pH = 12.1, ~~it~~ it will not function as a buffer.

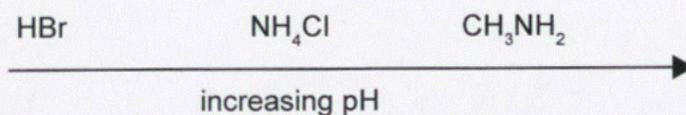
$\text{p}K_a$ is a constant and therefore can't be changed.

pH can be lowered by adding more acid because $\text{pH} = -\log[\text{H}_3\text{O}^+]$.

$\text{pH} = \text{p}K_a + \log\left[\frac{b}{a}\right]$ When the concentration of [base] and [acid] are equal then $\log\left[\frac{b}{a}\right]$ will = 0 and therefore $\text{pH} = \text{p}K_a$.

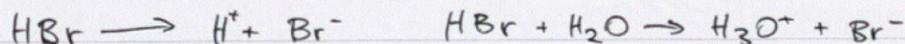
This is when the buffer would be most effective.

- (c) The pH of three solutions of equal concentration were ranked in order of increasing pH.



Justify the order in terms of the degree of dissociation and the relative concentration of hydronium ions in each solution.

Include relevant equation(s) in your answer.



pH is relative to the $[\text{H}_3\text{O}^+]$. $\text{pH} = -\log [\text{H}_3\text{O}^+]$

the higher the $[\text{H}_3\text{O}^+]$, the lower the pH.

HBr is a strong acid and fully dissociates into its ions giving a high $[\text{H}_3\text{O}^+]$ and therefore the lowest pH of the three solutions.



NH_4Cl is a salt and first fully dissociates into NH_4^+ and Cl^- .

The NH_4^+ ^(weak acid) reacts with water and partially dissociates into NH_3 and H_3O^+ .

$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$. This means a lower $[\text{H}_3\text{O}^+]$ than HBr but still a higher $[\text{H}_3\text{O}^+]$ than $[\text{OH}^-]$, giving a pH ~~lower~~ higher than HBr but still lower than 7.



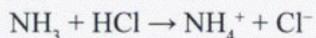
CH_3NH_2 is a weak base and partially dissociates into its ions,

CH_3NH_3^+ and OH^- . $2\text{H}_2\text{O} \rightleftharpoons \text{OH}^- + \text{H}_3\text{O}^+$. - in this equilibrium it shows that water still produces H_3O^+ but a very low concentration and a lower concentration than $[\text{OH}^-]$ therefore giving a higher pH than both HBr and NH_4Cl .

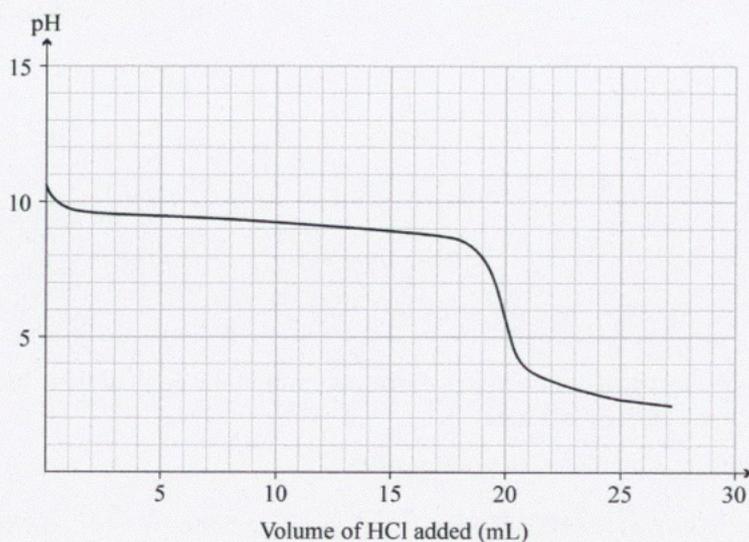
QUESTION THREE

A titration was carried out by adding $0.0174 \text{ mol L}^{-1}$ hydrochloric acid, $\text{HCl}(aq)$, to 25.0 mL of $0.0139 \text{ mol L}^{-1}$ ammonia, $\text{NH}_3(aq)$, in a conical flask.

The equation for the reaction is:



$$K_a(\text{NH}_4^+) = 5.75 \times 10^{-10} \quad \text{p}K_a(\text{NH}_4^+) = 9.24$$



- (a) Explain, with reference to the titration curve, why methyl red ($\text{p}K_a$ 5.1) would be a better indicator to detect the equivalence point than thymol blue ($\text{p}K_a$ 1.7).

This titration is between a strong acid and a weak base. To know when all the reactants have formed products an indicator can be used but it must change colour at or around the equivalence point to be effective.

At $\text{p}K_a$ 1.7 Using thymol blue would not be effective because the pH wouldn't go low enough for the indicator to have a colour change. methyl red would effectively show the point where all reactions have gone to completion with a $\text{p}K_a$ of 5.1 and the equivalence point being slightly above 5.

(b) (i) Calculate the pH at equivalence point.

initial guess: approx 5.5

$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$



$$\begin{aligned} \text{pH} &= \text{p}K_a + \log \left[\frac{b}{a} \right] \\ &= 9.24 + \log [] \end{aligned}$$

assume $[\text{NH}_3] = [\text{H}_3\text{O}^+]$

$$K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{NH}_4^+]}$$

$$[\text{NH}_4^+] = 0.0174 \times \frac{20}{45} = 7.73 \times 10^{-3} \text{ mol L}^{-1}$$

$$[\text{H}_3\text{O}^+] = \sqrt{K_a \times [\text{NH}_4^+]}$$

$$[\text{H}_3\text{O}^+] = \sqrt{(5.75 \times 10^{-10}) \times (7.73 \times 10^{-3})}$$

$$[\text{H}_3\text{O}^+] = 2.11 \times 10^{-6} \text{ mol L}^{-1} \text{ (3sf)}$$

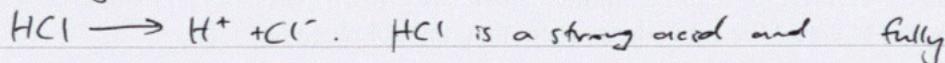
$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (2.11 \times 10^{-6})$$

$$\text{pH} = 5.676 \quad \text{pH} = 5.68 \text{ (3sf)}$$

(ii) Compare the electrical conductivity of the solution in the conical flask before HCl is added, and after 20 mL of HCl is added.

Include relevant equation(s) in your answer.

electrical conductivity depends on the amount of ions present that can carry charge.

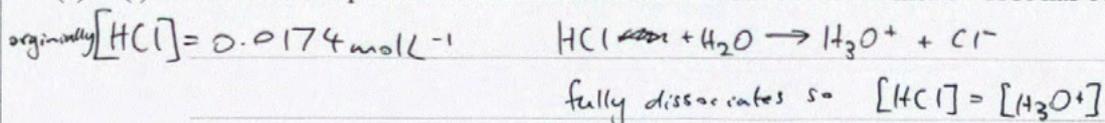


dissociates into ion giving a high concentration of ions that can carry charge. Therefore adding HCl makes the solution

a better electrical conductor than before HCl was added and it was just $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$

a small concentration of ions since NH_3 is a weak base and only partially dissociates into its ions.

- (c) (i) Calculate the pH in the conical flask after 25.0 mL of 0.0174 mol L⁻¹ HCl has been added.



$$[H_3O^+] = 0.0174 \times \frac{5}{25+25} = 1.74 \times 10^{-3} \text{ mol L}^{-1}$$

$$pH = -\log [H_3O^+] = -\log(1.74 \times 10^{-3}) = 2.759$$

$$pH = 2.76 \text{ (3sf)}$$

- (ii) The 0.0174 mol L⁻¹ HCl in the burette has a pH of 1.76.

Explain why the pH calculated in part (i) is higher than 1.76.

No calculations are necessary.

The pH is slightly higher than the original pH of HCl because ~~the~~ base ~~NH₃~~ is being added to ^{the} base, NH₃.

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$. The increase in $[OH^-]$ means OH^- is neutralising H_3O^+ . $OH^- + H_3O^+ \rightleftharpoons 2H_2O$. This decreases

$[H_3O^+]$. $pH = -\log [H_3O^+]$ and therefore pH will be slightly higher.

Extra space if required.
Write the question number(s) if applicable.

QUESTION
NUMBER

$$\text{pH} = \text{pK}_a + \log \left[\frac{\text{b}}{\text{a}} \right]$$

Extra space if required.
Write the question number(s) if applicable.

QUESTION
NUMBER

91392

Merit

Subject: L3 Chemistry

Standard: 91392

Total score: 16

Q	Grade score	Marker commentary
One	M5	<p>The candidate was awarded M5 for the following reasons:</p> <p>In part (a), the candidate correctly calculated the solubility product. Furthermore, the candidate explained that a decrease in $[\text{OH}^-]$ due to reaction with H_3O^+ will favour formation of more OH^- in the saturated solution of iron(II) hydroxide.</p> <p>In part (b), the candidate explained that the formation of the complex ion causes a decrease in $[\text{Ni}^{2+}]$ and subsequently favours the forward reaction.</p> <p>The candidate calculated an incorrect ionic product, and this has prevented the candidate from gaining Excellence for Question One.</p>
Two	M6	<p>The candidate was awarded M6 for the following reasons:</p> <p>In part (b), the candidate explained how the buffer resists a change in pH when NaOH is added. Furthermore, the candidate calculated the correct mass of methylammonium chloride required to prepare a solution at a pH of 12.1. The candidate then outlined why the solution cannot function as a buffer and how it could be modified to function as a buffer. However, the candidate did not clearly specify how the solution could be modified in the laboratory by adding either a strong acid or more solid methylammonium chloride. This has prevented the candidate from gaining Excellence for Question Two.</p> <p>In part (c), the candidate justified the order in pH in terms of the degree of dissociation and $[\text{H}_3\text{O}^+]$, supported by relevant equations.</p>
Three	M5	<p>The candidate was awarded M5 for the following reasons:</p> <p>In part (b), the candidate correctly calculated the pH at the equivalence point. In addition, the candidate linked the electrical conductivity of the ammonia solution to the degree of dissociation and [ions] in solution.</p> <p>In part (c), the candidate correctly calculated the pH after the equivalence point, once 25.0mL of HCl had been added.</p> <p>To gain M6, the candidate needed to explain for part (a) that the thymol blue would be an unsuitable indicator, since its pH range of 0.7 – 2.7 would mean the indicator would change colour after the equivalence point, whereas methyl red changes colour over a pH range of 4.1 – 6.1 and would therefore change colour at the estimated equivalence point.</p>