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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.

Achievement

TOTAL 12

QUESTION ONE

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

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

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

$K_s = [Fe^{2+}][OH^-]^2$

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

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

$K_s = 4(1.01 \times 10^{-5})^3$
$= 4.12 \times 10^{-15}$ mol L^{-1}

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

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

Include relevant equation(s) in your answer.

No calculations are necessary.

HCl is a strong acid which will lower the pH of the solution. lower pH means less OH^-. this will favour the forward reaction ~~to~~ to increase production of OH^- to make up for loss. this will use up more Fe(OH) meaning more will disolve and solubility increases.

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

(i) A student makes a green precipitate of $Ni(OH)_2$ in a test tube by mixing sodium hydroxide, $NaOH(aq)$, and nickel chloride, $NiCl_2(aq)$. When the student then adds excess sodium cyanide, $NaCN(aq)$, the green precipitate dissolves.

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

by adding NaCn to $Ni(OH)_2$ in solution Ni and Cn formed the complex Ion $[Ni(CN)_4]^{2-}$ shown by:

$Ni^{2-} + 4CN^{+} \longrightarrow [Ni(CN)_4]^{2-}$

(ii) Determine whether a precipitate of $Ni(OH)_2$ 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.

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

$[Ni] = 0.130 \times 55/90$
$= 0.079$

$NiSO_4$ → Ni^{2-}, SO_4^{4}

KOH → K^+, OH^-

$Ip = [Ni][OH]^2$
$= [0.079][1.95 \times 10^{-3}]^2$
$= 3.018 \times 10^{-7}$

$Ip > K_s$ so a precipitate will form.

$pOH = 14 - 11.7$
$= 2.3$

$[OH] = \cancel{-\log 2.3}$
$10^{-2.3}$
$= 5.012 \times 10^{-3}$

$\times 35/90 = 1.95 \times 10^{-3}$ molL⁻

QUESTION TWO

Methylammonium chloride, CH_3NH_3Cl, is an acidic salt.

(a) List all the species present in a solution of CH_3NH_3Cl in order of decreasing concentration. Do not include water.

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

Space for working if required

CH_3NH_3Cl ~~$+H_2O$~~ $\longrightarrow CH_3NH_3^+ + Cl^-$

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

(b) If CH_3NH_3Cl and methanamine, CH_3NH_2, are mixed in the appropriate quantities, a buffer solution is formed.

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

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

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

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 reduce pH change they do this by utilizing a combination of a weak base and its conjugate acid. When a strong base is added the OH^- component is removed/neutralized by the weak acid / $CH_3NH_3^+$ and turned into weak base. The created weak base CH_3NH_2 is much weaker than NaOH so pH change is limited.

(ii) Calculate the mass of CH_3NH_3Cl that should be added to 600 mL of 0.840 mol L^{-1} CH_3NH_2 to produce a solution of pH 12.1.

$M(CH_3NH_3Cl) = 67.5$ g mol^{-1}

$\cancel{12}$ $pH = pK_a + \log \frac{b}{a}$

$12.1 = 10.6 + \log \frac{b}{a}$

$\cancel{Rb}$ $n_b = 0.84 \times 0.6$

$= .504$

$\cancel{m_b = .504 \times 67.5}$

$\cancel{= 34.02}$

$-\log 12.1 = -\log 10.6 + \frac{.504}{a}$

$-1.08 = -1.02 + \frac{.504}{a}$

$-2.108 = -\frac{.504}{a}$

$a = \frac{-.504}{-2.108}$

$n_a = .23$

$m_a = .23 \times 67.5$

$= 16.14$ g

(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.

as pH equal 12.1 it is outside buffer zone of pKa (10.6) $\pm$ 1. Instead the Solution will crash. by adding a higher concentration of weak acid ~~or~~ more of it pH will come down within the buffer zone of 10.6 $\pm$ 1

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

HBr $\quad$ NH_4Cl $\quad$ CH_3NH_2

$\xrightarrow{\text{increasing pH}}$

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.

HBr is a relatively strong acid that ~~completely~~ ~~dissacociates~~ partially dissacociates

$HBr + H_2O \rightleftharpoons Br^- + H_3O^+$

NH_4Cl is an acidic salt that also partially dissacociates to make H_3O^+

NH_4Cl ~~$+H_2O$~~ $\longrightarrow NH_4^+ + Cl^-$

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

This dissacociation is to a lesser degree then HBr so it creates a lower concentration ~~at~~ of H_3O^+ and has a higher pH

CH_3NH_2 is a weak base that partially dissacociates to create OH^- and an extremely low conc of H_3O^+ making it have the highest pH

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

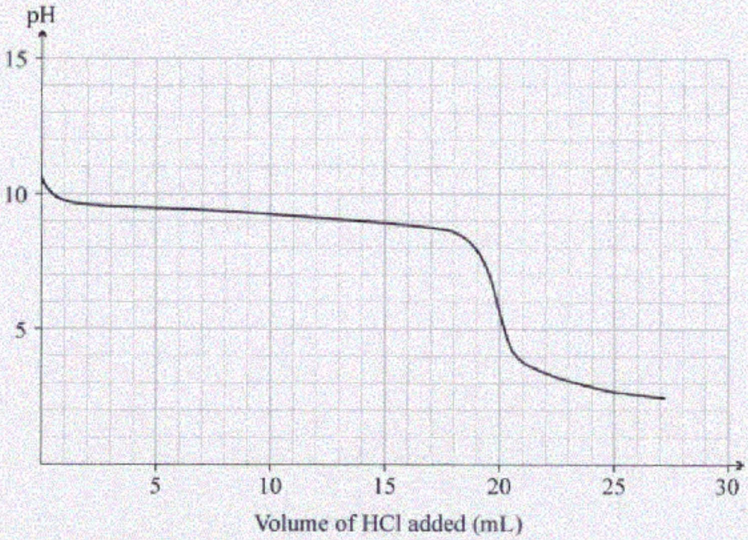
QUESTION THREE

A titration was carried out by adding 0.0174 mol L^{-1} hydrochloric acid, $HCl(aq)$, to 25.0 mL of 0.0139 mol L^{-1} ammonia, $NH_3(aq)$, in a conical flask.

The equation for the reaction is:

$NH_3 + HCl \rightarrow NH_4^+ + Cl^-$

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



(a) Explain, with reference to the titration curve, why methyl red (pK_a 5.1) would be a better indicator to detect the equivalence point than thymol blue (pK_a 1.7).

equivalence point is at ≈ 5–6 pH. a good indicator to detect this would be ± pka at this point thymol blue is effective at a pH of 1.7 which is too low. methyl red is effective at 5.1 which is within range meaning it would detect the equivalence point

INTERUPTION

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

~~pH = pKa Ka = [H₃O⁺]~~

$$[H_3O^+] = \sqrt{\frac{Ka \times Kw}{[base]}}$$

$$= \sqrt{\frac{5.75 \times 10^{-10} \times 1 \times 10^{-14}}{0.0137 \times 25/45}}$$

$$= 2.728 \times 10^{-11}$$

$$pH = -\log 2.728 \times 10^{-11}$$

$$= 10.56$$

(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.

before HCl is added NH_3 in Solution partially dissacociates to create a lot of free moving Ions to carry charge making it reletively conductive shown by

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

at equivalenc point

$NH_3 + HCl + H_2O \rightleftharpoons NH_4^+ + OH^- + Cl^- + H_3O^+$

more free moving Ions to carry charge so more conductive

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

(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.

because the HCl is diluted as it is added to the 25 ml solution raising its pH.

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Write the question number(s) if applicable.

QUESTION NUMBER

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QUESTION NUMBER

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Achievement

Subject: L3 Chemistry

Standard: 91392

Total score: 12

Q	Grade score	Marker commentary
One	A4	The candidate was awarded A4 for the following reasons: In part (a), the candidate wrote the equilibrium equation and K_s expression, and calculated the solubility product. In part (b), the candidate recognised that the addition of cyanide ions will form a complex ion with Ni^{2+} ions. Although individual ions have the incorrect charges, the complex ion is correct. Furthermore, the candidate correctly calculated the ionic product and compared with the solubility product to identify that a precipitate of nickel hydroxide would form.
Two	A4	The candidate was awarded A4 for the following reasons: In part (a), the candidate listed the correct remaining species present in the solution of methylammonium chloride. In part (b), the candidate explained how the buffer resists a change in pH when NaOH is added. In addition, the candidate explained why the solution could not function as a buffer and how it could be modified to function as a buffer. In part (c), the candidate recognised that CH_3NH_2 has the lowest $[H_3O^+]$.
Three	A4	The candidate was awarded A4 for the following reasons: In part (a), the candidate wrote the equilibrium equation and K_s expression, and calculated the solubility product. In part (b), the candidate recognised that the addition of cyanide ions will form a complex ion with Ni^{2+} ions. Although individual ions have the incorrect charges, the complex ion is correct. Furthermore, the candidate correctly calculated the ionic product and compared with the solubility product to identify that a precipitate of nickel hydroxide would form.