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91392



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Level 3 Chemistry, 2016

91392 Demonstrate understanding of equilibrium principles in aqueous systems

2.00 p.m. Monday 21 November 2016
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 is provided in the Resource Sheet L3-CHEMR.

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

Check that this booklet has pages 2–8 in the correct order and that none of these pages is blank.

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

TOTAL

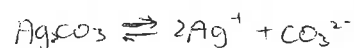
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QUESTION ONE

Silver carbonate, Ag_2CO_3 , is a sparingly soluble salt.



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

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

$$K_s(\text{Ag}_2\text{CO}_3) = [\text{Ag}^+(\text{aq})]^2 [\text{CO}_3^{2-}(\text{aq})]$$

- (b) Calculate the mass of Ag_2CO_3 that will dissolve in 50 mL of water to make a saturated solution at 25°C .

$$K_s(\text{Ag}_2\text{CO}_3) = [\text{Ag}^+(\text{aq})]^2 [\text{CO}_3^{2-}(\text{aq})]$$



let $[\text{Ag}^+]$ be $2s$, $[\text{CO}_3^{2-}]$ be s .

$$8.10 \times 10^{-12} = (2s)^2 \cdot s$$

$$4s^3 = 8.10 \times 10^{-12}$$

$$s = 1.27 \times 10^{-4} \text{ mol L}^{-1} \text{ (3sf)}$$

$$\therefore [\text{Ag}_2\text{CO}_3] = s = 1.27 \times 10^{-4} \text{ mol L}^{-1}$$

$$n = cV = 1.27 \times 10^{-4} \times 0.0500 = 6.33 \times 10^{-6} \text{ mol}$$

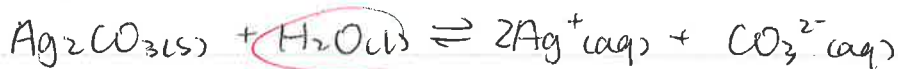
$$m = nM = 6.33 \times 10^{-6} \times 276$$

$$= 1.75 \times 10^{-3} \text{ g (3sf)}$$

- (c) Explain how the solubility of Ag_2CO_3 will change if added to $\overset{V}{50 \text{ mL}}$ of a $\overset{c}{1.00 \text{ mol L}^{-1}}$ ammonia, NH_3 , solution.

Support your answer with balanced equations.

No calculations are necessary.



The 1.00 mol L^{-1} ammonia solution is concentrated enough to complex Ag^+ out. $\text{Ag}^+ + 2\text{NH}_3 \rightarrow [\text{Ag}(\text{NH}_3)_2]^+$ (Tollen Reagent). In this way, Ag^+ is removed and to minimise the change, the equilibrium system will shift to the right and now products are favoured. Therefore, more Ag_2CO_3 will dissolve and the solubility will increase.

- (d) Show by calculation whether a precipitate of Ag_2CO_3 will form when $\overset{V}{20.0 \text{ mL}}$ of $\overset{c}{0.105 \text{ mol L}^{-1}}$ silver nitrate, AgNO_3 , solution is added to $\overset{V}{35.0 \text{ mL}}$ of a $\overset{c}{0.221 \text{ mol L}^{-1}}$ sodium carbonate, Na_2CO_3 , solution.

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

$$n(\text{AgNO}_3) = cV = 0.020 \times 0.105 = 2.10 \times 10^{-3} \text{ mol}$$

$$\text{new concentration } [\text{Ag}^+] = [\text{AgNO}_3] = \frac{n}{V} = \frac{2.10 \times 10^{-3}}{0.055} = 0.0382 \text{ mol L}^{-1}$$

$$n(\text{Na}_2\text{CO}_3) = cV = 0.035 \times 0.221 = 7.74 \times 10^{-3} \text{ mol}$$

$$\text{new } [\text{CO}_3^{2-}] = [\text{Na}_2\text{CO}_3] = \frac{n}{V} = \frac{7.74 \times 10^{-3}}{0.055} = 0.141 \text{ mol L}^{-1}$$



$$IP(\text{Ag}_2\text{CO}_3) = [\text{Ag}^+]^2 [\text{CO}_3^{2-}]$$

$$= (0.0382)^2 \times 0.141$$

$$= 2.05 \times 10^{-4} \text{ mol L}^{-1}$$

$$\therefore 2.05 \times 10^{-4} > 8.10 \times 10^{-12}, IP > K_s$$

\therefore A precipitate will form.

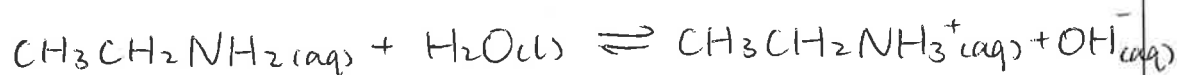
QUESTION TWO

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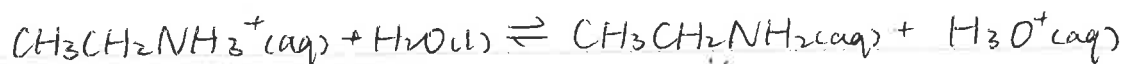
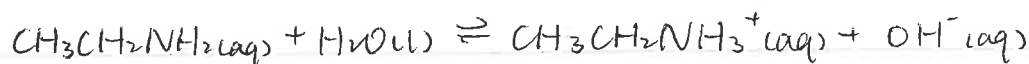
Ethanamine, $\text{CH}_3\text{CH}_2\text{NH}_2$, is a weak base.

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

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



(b) Calculate the pH of a 0.109 mol L^{-1} solution of ethanamine.



$$\text{Assume } [\text{OH}^-] = [\text{CH}_3\text{CH}_2\text{NH}_3^+] = \frac{K_w}{[\text{H}_3\text{O}^+]}$$

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

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

$$\therefore 2.51 \times 10^{-11} = \frac{[\text{H}_3\text{O}^+]^2 \times 0.109}{1 \times 10^{-14}}$$

$$[\text{H}_3\text{O}^+]^2 = 2.30 \times 10^{-24}$$

$$[\text{H}_3\text{O}^+] = 1.52 \times 10^{-12} \text{ mol L}^{-1} \text{ (3sf)}$$

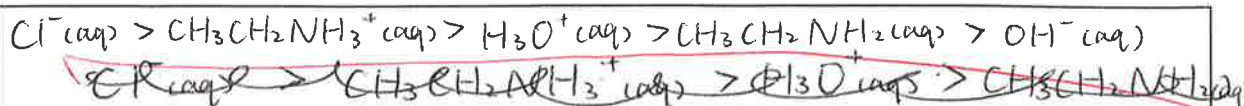
$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (1.52 \times 10^{-12}) = 11.8 \text{ (3sf)}$$

- (c) Ethyl ammonium chloride, $\text{CH}_3\text{CH}_2\text{NH}_3\text{Cl}$, is a weak acid that will also react with water.

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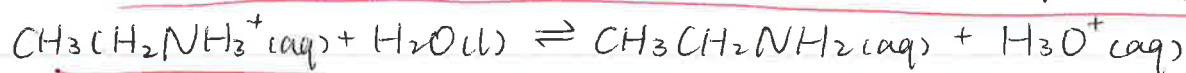
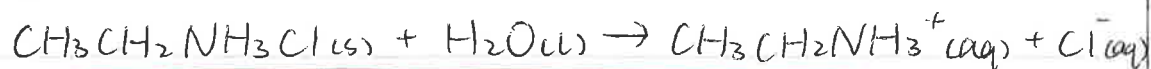
List all the species present in a solution of $\text{CH}_3\text{CH}_2\text{NH}_3\text{Cl}$, in order of decreasing concentration.

Do not include water.



Justify the order you have given.

Include equations, where necessary.



$\text{CH}_3\text{CH}_2\text{NH}_3\text{Cl}$ is an ionic compound so it fully dissociates in water and produces Cl^- and $\text{CH}_3\text{CH}_2\text{NH}_3^+$. Cl^- does not have further reaction so is in the highest concentration. $\text{CH}_3\text{CH}_2\text{NH}_3^+$ has further reaction with water as it is ^{a weak acid} ~~the conjugate~~. It only partially dissociates in water producing H_3O^+ and $\text{CH}_3\text{CH}_2\text{NH}_2$. Only a small ^{proportion} ~~portion~~ reacts so the concentration of $\text{CH}_3\text{CH}_2\text{NH}_3^+$ is the second highest. Although we assume $\text{CH}_3\text{CH}_2\text{NH}_2$ and H_3O^+ are in equal concentration, there is a small contribution to H_3O^+ from water so $[\text{H}_3\text{O}^+]$ is slightly larger than $[\text{CH}_3\text{CH}_2\text{NH}_2]$. Lastly, there are OH^- ions in all solutions so OH^- exists but in a very small concentration (acidic solution, $[\text{H}_3\text{O}^+] > [\text{OH}^-]$).

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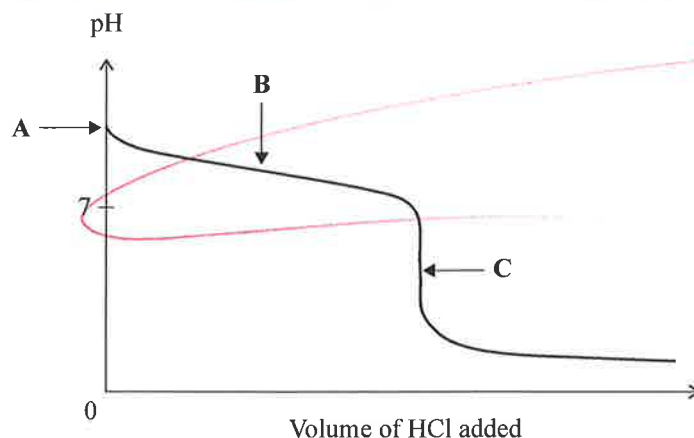
QUESTION THREE

20.00 mL of 0.320 mol L⁻¹ ammonia, NH₃, is titrated with 0.640 mol L⁻¹ hydrochloric acid, HCl.

The equation for this reaction is:



The curve for this titration is given below.



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

At the equivalence point, equal moles of acid and base has reacted and the salt (NH_4^+ and Cl^-) are in the highest concentration. However, NH_4^+ further react with water. It is a weak acid so it partially dissociates: $\text{NH}_4^+(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$. The reaction produces H_3O^+ so now $[\text{H}_3\text{O}^+] > [\text{OH}^-]$ and the pH is not 7 but lower than 7.

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

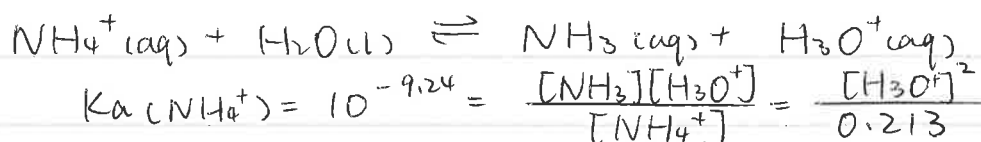
$$C_1 V_1 = C_2 V_2$$

$$0.320 \times 0.020 = \frac{0.640 \times V_2}{\cancel{[\text{NH}_4^+]}} \times \cancel{(0.020 + 0)}$$

$$V_2 = 0.0100 \text{ L}$$

$$0.320 \times 0.020 = (0.020 + 0.010) [\text{NH}_4^+]$$

$$[\text{NH}_4^+] = 0.213 \text{ mol L}^{-1}$$



$$[\text{H}_3\text{O}^+] = \sqrt{1.11 \times 10^{-5}} \text{ mol L}^{-1}$$

Chemistry 91392, 2016 $\text{pH} = -\log [\text{H}_3\text{O}^+]$

$$= 4.955471 = 4.96 (3 \text{ sf})$$

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

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At this point, NH_4^+ and NH_3 are present in approximately equal concentration. $K_a(\text{NH}_4^+) = \frac{[\text{H}_3\text{O}^+][\text{NH}_3]}{[\text{NH}_4^+]} = [\text{H}_3\text{O}^+]$ at this point ($[\text{NH}_4^+] = [\text{NH}_3]$). So $\text{p}K_a = \text{pH} = 9.24$ at this point.

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

No calculations are necessary.

At point C, equal moles of acid and base has reacted. Cl^- is in the highest concentration as it does not further react. NH_4^+ is a weak acid so it has further reaction with water, where it partially dissociates into NH_3 and H_3O^+ . In this way, $[\text{H}_3\text{O}^+]$ is larger than $[\text{OH}^-]$ and the pH is approximately half of the $\text{p}K_a(\text{NH}_4^+)$ value ($9.24/2 = 4.62$).

E8

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

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QUESTION
NUMBER

91392

Excellence exemplar 2016

Subject:	Chemistry	Standard:	91392	Total score:	22
Q	Grade score	Annotation			
1	E7	<p>The candidate has completed the solubility product expression, K_s, correctly in part (a).</p> <p>In part (b), the correct methods are used to calculate the solubility, the number of moles and the correct mass.</p> <p>A very good explanation is given for part (c), however, the inclusion of H_2O within the equation prevented a higher grade score from being attained.</p> <p>In part (d), the calculation is fully correct, however, I.P. does not have the units $mol\ L^{-1}$.</p>			
2	E7	<p>The candidate gave the correct reaction for ethanamine with water in part (a), and calculated the pH of the solution of ethanamine correctly.</p> <p>In part (c), the candidate has given the correct species and a good discussion as to why they form in the quantities that they do. A fuller justification is required as to the reason for the formation of hydroxide ions in lowest concentration.</p>			
3	E8	<p>Full credit is given for all question parts.</p> <p>The pH is correctly explained in terms of the species present, supported with appropriate equations.</p> <p>The concentration of the ammonium ions is correctly calculated and used to justify the pH.</p> <p>The pH is explained in terms of both the species present, and by mathematical relationship.</p>			