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91392

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NZQA

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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.

Merit

TOTAL 17

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QUESTION ONE

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Silver carbonate, Ag_2CO_3, is a sparingly soluble salt.

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

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

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

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

$m = n/M$ $n = cV$

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

$2x$ x

$K_s = 4x^3$

$\sqrt[3]{\frac{8.10 \times 10^{-12}}{4}} = \sqrt[3]{\frac{4x^3}{4}}$

$x = 1.265 \times 10^{-4}$ mol L^{-1}

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

$n = 1.27 \times 10^{-4} \times (50/1000)$

$= 6.326 \times 10^{-6}$ mol

$m = 6.326 \times 10^{-6} / 276$

$= 2.29 \times 10^{-8}$ g. (3sf)

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(c) 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.

No calculations are necessary.

The solubility will ~~decrease~~ increase, as there will be more solids ~~forming~~ will be soluble. The NH_3 will ~~react~~ react with CO_3^{2-}. Because there is a decrease of CO_3^{2-}, then the equilibrium will favour the forward reaction to compensate for the loss CO_3^{2-}. This would result in a more soluable solution.

(d) 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 25°C

$[Ag^+] = \frac{20}{55} \times 0.105 = 0.0382 \text{ mol } L^{-1}$

$[CO_3^{2-}] = \frac{35}{55} \times 0.221 = 0.141 \text{ mol } L^{-1}$

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

$= (0.0382)^2(0.141)$

$= 2.05 \times 10^{-4}$ (3sf).

$Q > K_s$ ∴ precipitate forms.

MS

QUESTION TWO

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Ethanamine, $CH_3CH_2NH_2$, is a weak base.

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

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

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

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

$$K_b = \frac{[CH_3CH_2NH_3^+][OH^-]}{[CH_3CH_2NH_2]}$$

$K_b = \frac{1 \times 10^{-14}}{2.51 \times 10^{-11}}$

$= 3.98 \times 10^{-4}$

$3.98 \times 10^{-4} = \frac{[CH_3CH_2NH_3^+][OH^-]}{(0.109)}$

Assume $[CH_3CH_2NH_3^+] = [OH^-]$.

$3.98 \times 10^{-4} = \frac{[OH^-]^2}{0.109}$

$\sqrt{(3.98 \times 10^{-4})(0.109)} = \sqrt{[OH^-]^2}$

$[OH^-] = 6.59 \times 10^{-3}$

$pOH = -\log [OH^-]$

$= -\log (6.59 \times 10^{-3})$

$= 2.18$

$pKb = 14 - 10.6$

$= 3.4$

$pOH = 10^{-3.4}$

$= 3.98 \times 10^{-4}$

$[OH^-] = 10^{-3.98 \times 10^{-4}}$

$= 1.00$

$pH = 10^{-10.6}$

$= 2.51 \times 10^{-11}$

$pH + pOH = 14$

$pH = 14 - 2.18$

$= 11.82.$

$= 11.8$ (3sf)

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

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

(c) Ethyl ammonium chloride, $CH_3CH_2NH_3Cl$, is a weak acid that will also react with water.

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List all the species present in a solution of $CH_3CH_2NH_3Cl$, in order of decreasing concentration.

Do not include water.

$[Cl^-] > [CH_3CH_2NH_3^+] > [H_3O^+] \geq [CH_3CH_2NH_2] \gg [OH^-]$

Justify the order you have given.

Include equations, where necessary.

$[Cl^-]$ is the largest species present due to it not being used for further reactions. $[CH_3CH_2NH_3^+]$ is the second largest species present due to the ~~formation~~ splitting of the H^+ from $CH_3CH_2NH_3Cl$, but it is used for further reaction with water, thus it is smaller than $[Cl^-]$. $[H_3O^+]$ is slightly larger ~~than~~ than ~~$[CH_3CH_2NH_2]$~~ or is equal to $[CH_3CH_2NH_2]$ due to the further reaction with water producing its products. $[OH^-]$ is ~~small~~ the smallest because, there isn't much OH^- ions producing as $CH_3CH_2NH_3Cl$ is a weak acid, which means that there is more $[H_3O^+]$ present than $[OH^-]$, thus it is the smallest species present.

E7

QUESTION THREE

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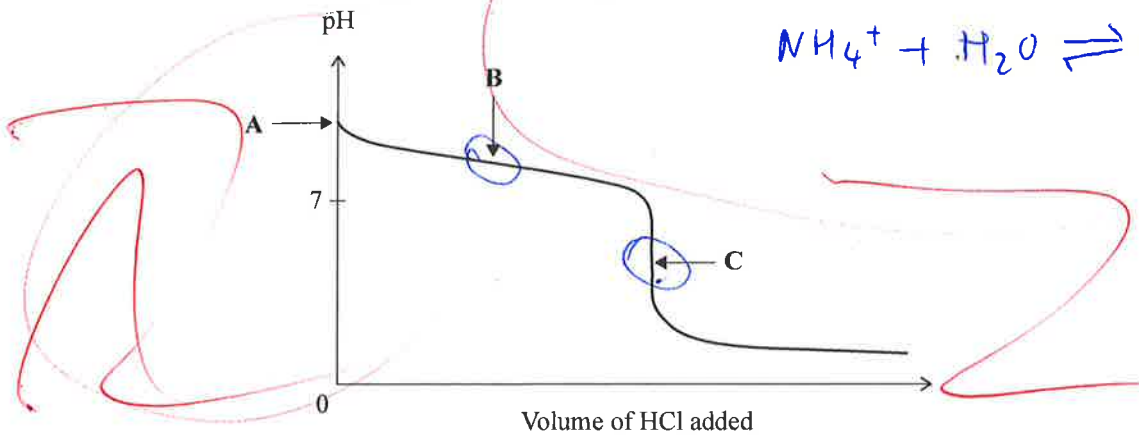
20.00 mL of 0.320 mol L^{-1} ammonia, NH_3, is titrated with 0.640 mol L^{-1} hydrochloric acid, HCl.

The equation for this reaction is:

$NH_3 + HCl \rightarrow NH_4^+ + Cl^-$ $\quad pK_a(NH_4^+) = 9.24$ $\quad Ka = -0.966$

The curve for this titration is given below.



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

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

At Point C, equivalence point is not 7 due to all the species being neutralised however, there are some unreacted HCl, thus the pH will be lower than 7 as HCl is a strong acid, with more $[H_3O^+]$ ions than $[OH^-]$, thus point C is lower than 7.

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

$pKa = 9.24$

$pH = 10^{-pKa}$

$pH = pKa + \left(\log \frac{\text{base}}{\text{acid}}\right)$ at C

$[NH_3] = [H_3O]$

$=$

$Ka = \frac{[H_3O^+]^2}{(0.320)}$

$Ka = 3.76 \times 10^{-10}$

(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 point B, is a buffer zone, This means that half the species are reacted. Because half the species are reacted, there would be minimal movement of pH, thus pKa ≈ pH, so 9.24 ≈ 9.24

half of NH_3 is reacted with HCl

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

No calculations are necessary.

The pH of the solution at point C is 4.96. This is because, at this point all the species present (NH_3) has been neutralised with (HCl). However, the pH is a weak acid (NH_4^+) because there are unreacted or excess H_3O^+ it the solution, which cause the pH to lower.

M5

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

QUESTION NUMBER

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91392

Merit exemplar 2016

Subject:	Chemistry	Standard:	91392	Total score:	17

Q	Grade score	Annotation
1	M5	The candidate has completed the solubility product expression, K_s, correctly in part (a). In part (b), the correct methods are used to calculate the solubility and number of moles, but an incorrect equation is used to calculate the final step of the calculation. No equations were included in the explanation for part (c). Both dilutions in part (d) are correctly calculated. K_s is also correctly calculated and a suitable conclusion is given based on this calculation.
2	E7	The candidate gave the correct reaction for ethanamine with water in part (a), and calculated the pH of the solution of ethanamine correctly. 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.
3	M5	The candidate's response for part (a) explained the level of acidity. No relevant evidence is provided in part (b). No evidence of the species present or mathematical relationship is provided to explain why the pH at point B on the graph is 9.24. In part (d), the candidate outlines that the pH is that of a weak acid due to the presence of the ammonium ion. The species is linked to acidity and to an equation.