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# 3

91392



913920



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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.**

**Achievement**

**TOTAL**

**11**

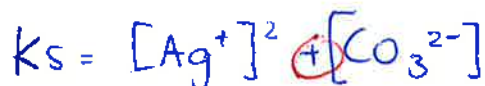
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## QUESTION ONE

Silver carbonate,  $\text{Ag}_2\text{CO}_3$ , is a sparingly soluble salt.

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

- (a) Write the solubility product expression,  $K_s$ , for silver carbonate ( $\text{Ag}_2\text{CO}_3$ ).



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

$$\frac{1000}{50} = 20$$

$$K_s = 4s^3$$

$$\sqrt[3]{\frac{4s}{4}} = s$$

$$\sqrt[3]{\frac{8.10 \times 10^{-12}}{4}} = s = 1.265148998 \times 10^{-4} \text{ mol L}^{-1}$$

$$20 \times \frac{1.265 \dots \times 10^{-4}}{\text{mol}} = \text{in } 50 \text{ mL} = 6.32574499 \times 10^{-6}$$

$$2.530297996 \times 10^{-3}$$

$$n = 6.325 \dots \times 10^{-6}$$

$$n = \frac{m}{M_r}$$

$$6.325 \dots \times 10^{-6} \times 276$$

$$2.53029766 \times 10^{-3} \times 276 = m = 0.69883622$$

$$= 0.698 \text{ grams.}$$

- (c) Explain how the solubility of  $\text{Ag}_2\text{CO}_3$  will change if added to 50 mL of a 1.00 mol  $\text{L}^{-1}$  ammonia,  $\text{NH}_3$ , solution.

Support your answer with balanced equations.

No calculations are necessary.

~~$\text{Ag}_2\text{CO}_3 + \text{NH}_3 \rightleftharpoons$~~  Solubility will decrease because the addition of the base will interfere with the pH.

- (d) Show by calculation whether a precipitate of  $\text{Ag}_2\text{CO}_3$  will form when 20.0 mL of 0.105 mol  $\text{L}^{-1}$  silver nitrate,  $\text{AgNO}_3$ , solution is added to 35.0 mL of a 0.221 mol  $\text{L}^{-1}$  sodium carbonate,  $\text{Na}_2\text{CO}_3$ , solution.

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

$$0.105 = [\text{Ag}^+][\text{NO}_3^-]$$

$$[\text{Ag}^+] = \frac{0.105}{2} = 5.25 \times 10^{-2} \text{ mol L}^{-1}$$

$$= 2.1 \times 10^{-3}$$

$$0.221 = 2[\text{Na}^+]^2[\text{CO}_3^{2-}]$$

$$[\text{CO}_3^{2-}] = 7.735 \times 10^{-3}$$

$$K_s = \frac{[\text{Ag}^+]^2[\text{CO}_3^{2-}]}{[\text{Ag}_2\text{CO}_3]} = 3.41134 \times 10^{-8}$$

$\text{IP} < K_s$  so  
No precipitate.

It will not precipitate.

$$\frac{0.105}{50} = [\text{Ag}^+]$$

$$\frac{0.221}{\left(\frac{1000}{35}\right)} = [\text{CO}_3^{2-}]$$

## QUESTION TWO

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

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

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



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

$$K_a = \frac{[\text{NH}_3][\text{OH}^-]}{[\text{CH}_3\text{CH}_2\text{NH}_2]} \quad 2.51 \times 10^{-11} = \frac{[\text{OH}^-]^2}{0.109}$$

$$\sqrt{2.51 \times 10^{-11} \times 0.109} = [\text{OH}^-] = 1.6540556 \times 10^{-6}$$

$$-\log[\text{OH}^-] = 5.7814989$$

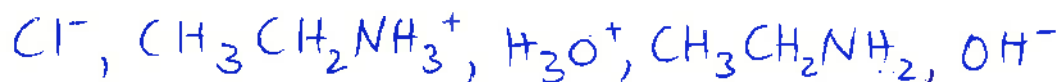
$$14 - 5.7814... = \text{pH} = 8.2185501$$

- (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.

$\text{CH}_3\text{CH}_2\text{NH}_3\text{Cl} \rightarrow \text{Cl}^- + \text{CH}_3\text{CH}_2\text{NH}_3^+$  This shows the complete dissociation of  $\text{CH}_3\text{CH}_2\text{NH}_3\text{Cl}$ . The  $\text{Cl}^-$  &  $\text{CH}_3\text{CH}_2\text{NH}_3^+$  are therefore the highest in concentration, however the  $\text{CH}_3\text{CH}_2\text{NH}_3^+$  undergoes an equilibrium reaction with  $\text{H}_2\text{O}$  (water) so some is used, making this concentration less than  $\text{Cl}^-$ .

$\text{CH}_3\text{CH}_2\text{NH}_3^+ + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{CH}_2\text{NH}_2 + \text{H}_3\text{O}^+$ . As this is an equilibrium reaction, the reactant is favoured most, so while product concentration is less. The  $\text{H}_3\text{O}^+$  reacts with excess water however producing both  $\text{H}_3\text{O}^+$  &  $\text{OH}^-$ , so the  $\text{H}_3\text{O}^+$  is of higher concentration than the  $\text{CH}_3\text{CH}_2\text{NH}_2$ .  $\text{OH}^-$  is the lowest as it only reacts is formed by the excess water in small quantities.

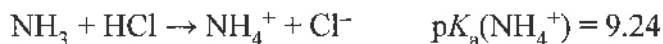
M5



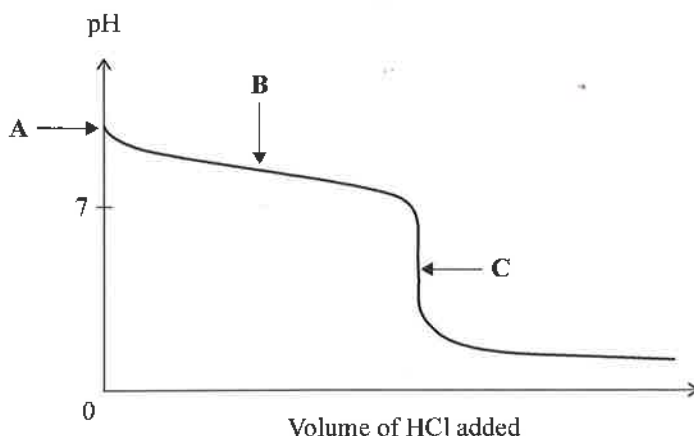
## QUESTION THREE

20.00 mL of 0.320 mol L<sup>-1</sup> ammonia, NH<sub>3</sub>, is titrated with 0.640 mol L<sup>-1</sup> 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 equivalence point, all of the NH<sub>3</sub> & HCl have reacted to form NH<sub>4</sub><sup>+</sup> & Cl<sup>-</sup>. The NH<sub>4</sub><sup>+</sup> product when reacted with water forms H<sub>3</sub>O<sup>+</sup>;  
 $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$  this H<sub>3</sub>O<sup>+</sup> is what contributes towards pH & so having a higher concentration means more contribution. This contribution is in favour of acid so the increased [H<sub>3</sub>O<sup>+</sup>] will decrease pH. The equivalence point

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

$$\text{p}K_a = 9.24 \quad K_a = \text{shift log}^{-1} \text{p}K_a = 5.754 \dots \times 10^{-10}$$

$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = 5.754399373 \times 10^{-10}$$

$$5.754 \dots \times 10^{-10} \times [\text{NH}_4^+] = [\text{H}_3\text{O}^+]^2$$

$$[\text{NH}_4^+] = \frac{0.32 + 0.64}{5} = 0.208929 \text{ mol L}^{-1}$$

$$\sqrt{5.754 \times 10^{-10} \times 0.2089 \dots} = [\text{H}_3\text{O}^+] = 1.096478196 \times 10^{-5}$$

$$-\log[\text{H}_3\text{O}^+] = \text{pH} = 4.9567$$

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

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

At halfway to the equivalence point, ~~the~~ point C, the concentration of  $\text{NH}_3$  is equal to its conjugate  $\text{NH}_4^+$ , so base concentration is equal to its conjugate acid. The effect this has is that  $\text{H}_3\text{O}^+$  concentration is the same as  $K_a$ , so  $\text{pH} = \text{p}K_a$ . ~~the~~  $\text{pH} = \text{p}K_a + \log \left[ \frac{\text{acid}}{\text{base}} \right]$   $\log \left[ \frac{\text{acid}}{\text{base}} \right] = 0$

pg 8. SEEN.

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

No calculations are necessary.

Because HCl is a strong acid, it will completely dissociate with the weak base  $\text{NH}_3$ . The  ~~$\text{H}_3\text{O}^+$~~   $\text{NH}_4^+$  &  $\text{Cl}^-$  formed,  $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NH}_3$  means that the pH will be acidic. The strong acid donates all of its  $\text{H}_3\text{O}^+$  towards the pH while the  $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$ ,  $\text{NH}_3$  only donates some of its  $\text{OH}^-$ . The pH is therefore very acidic, but not completely due to the  $\text{OH}^-$  being contributed.  $\cup$  1

AK

Extra paper if required.

Write the question number(s) if applicable.

QUESTION  
NUMBER

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3a. will therefore have a pH below 0, at the acidic <sup>side</sup> end of the scale.

3c. As the acid & base concentrations are equal therefore  $\text{pH} = \text{pKa}$  & as the pKa of the solution is 9.24, the pH will therefore be the same.

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Achievement exemplar 2016

Subject:	Chemistry	Standard:	91392	Total score:	11
Q	Grade score	Annotation			
1	N2	<p>The candidate incorrectly used a + sign in the solubility product expression, <math>K_s</math> in part (a).</p> <p>In part (b), the candidate correctly calculated the solubility, but has an incorrect number of moles. This error is carried through to give the final mass.</p> <p>Nothing of relevance is given in part (c).</p> <p>Both dilutions are incorrectly calculated in part (d), however, these values are carried through using the correct method to calculate <math>K_s</math>.</p>			
2	M5	<p>The candidate has given the wrong products in part (a), and has incorrectly calculated pOH in part (b).</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	A4	<p>This candidate made no reference to the concentration of hydronium ions being greater than the concentration of hydroxide ions in part (a).</p> <p>In part (b), the concentration of ammonium ions is incorrectly calculated.</p> <p>In part (c), the concepts were correctly used, however, for full marks the candidate needed to refer to a correct mathematical formula.</p>			