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# 3

91392



913920



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## Level 3 Chemistry, 2015

### 91392 Demonstrate understanding of equilibrium principles in aqueous systems

2.00 p.m. Wednesday 11 November 2015  
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 on 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 2–11 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.**

**Excellence**

TOTAL

**21**

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## QUESTION ONE

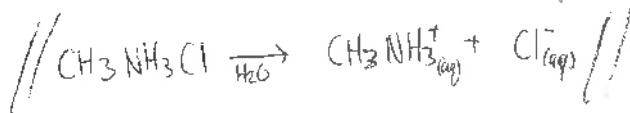
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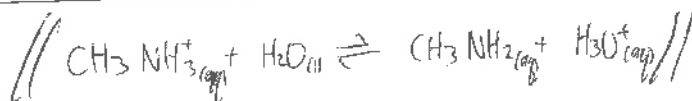
Methylammonium chloride,  $\text{CH}_3\text{NH}_3\text{Cl}$ , dissolves in water to form a weakly acidic solution.

$$K_a(\text{CH}_3\text{NH}_3^+) = 2.29 \times 10^{-11}$$

- (a) (i) Write an equation to show  $\text{CH}_3\text{NH}_3\text{Cl}$  dissolving in water.

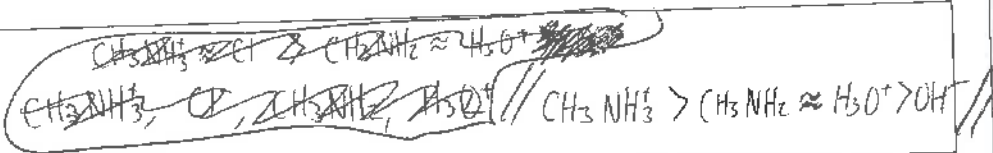


- (ii) Write an equation to show the reaction occurring in an aqueous solution of  $\text{CH}_3\text{NH}_3\text{Cl}$ .



- (iii) List all the species present in an aqueous solution of  $\text{CH}_3\text{NH}_3\text{Cl}$ , in order of decreasing concentration.

Do not include water.



- (iv) Calculate the pH of  $0.0152 \text{ mol L}^{-1}$   $\text{CH}_3\text{NH}_3\text{Cl}$  solution.

~~$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{NH}_2]}{[\text{CH}_3\text{NH}_3^+]}$~~  //  $K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{CH}_3\text{NH}_3^+]}$

$$2.29 \times 10^{-11} = \frac{[\text{H}_3\text{O}^+]^2}{0.0152}$$

$$[\text{H}_3\text{O}^+] = 5.90 \times 10^{-7} \text{ mol L}^{-1} \text{ (3s.f.)}$$

$$\text{pH} = -\log(5.90 \times 10^{-7})$$

$$= 6.23 \text{ (3s.f.)}$$

- (b) The table shows the pH and electrical conductivity of three solutions. The concentrations of the solutions are the same.

Solution	NaOH	CH <sub>3</sub> NH <sub>2</sub>	CH <sub>3</sub> COONa
pH	13.2	11.9	8.98
Electrical conductivity	good	poor	good

Compare and contrast the pH and electrical conductivity of these three solutions.

Include appropriate equations in your answer.

pH //  $\text{NaOH} \xrightarrow{\text{H}_2\text{O}} \text{Na}^+ + \text{OH}^-$  - NaOH fully dissociates in water because it is a strong base, to produce a high concentration of  $\text{Na}^+$  and  $\text{OH}^-$  ions. Because the concentration of  $\text{OH}^-$  ions produced is large, it has a high pH.

$\text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_3^+ + \text{OH}^-$  -  $\text{CH}_3\text{NH}_2$  is a weak base so only partially dissociates in water. The reaction  $\uparrow$  is reversible therefore there is only a small concentration of  $\text{OH}^-$  ions so it is only slightly basic (smaller pH than NaOH).

$\text{CH}_3\text{COONa} \xrightarrow{\text{H}_2\text{O}} \text{CH}_3\text{COO}^- + \text{Na}^+$ ,  $\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-$

-  $\text{CH}_3\text{COONa}$  fully dissociates in water to produce  $\text{CH}_3\text{COO}^-$  and  $\text{Na}^+$ .  $\text{CH}_3\text{COO}^-$  reacts with water to form  $\text{CH}_3\text{COOH}$  and  $\text{OH}^-$ , however the reaction is reversible because  $\text{CH}_3\text{COO}^-$  is a weak base, so there are only a small concentration of  $\text{OH}^-$  ions in solution, so the pH is lower than NaOH. //

// <sup>\*</sup>Electrical conductivity depends on the concentration of ions that are free to move and can carry charge

Electrical conductivity: \*  $\text{NaOH} \xrightarrow{\text{H}_2\text{O}} \text{Na}^+ + \text{OH}^-$  - NaOH fully dissociates in water to

produce a high concentration of  $\text{Na}^+$  and  $\text{OH}^-$  ions which are free to move and can carry charge, so NaOH is a good conductor of electricity.  $\text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_3^+ + \text{OH}^-$

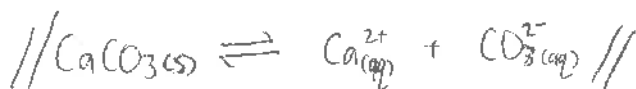
-  $\text{CH}_3\text{NH}_2$  only partially dissociates in water to produce a low concentration of  $\text{CH}_3\text{NH}_3^+$  and  $\text{OH}^-$  ions, so there are a smaller concentration of ions that are free to move and can carry charge, so  $\text{CH}_3\text{NH}_2$  is a poor conductor.  $\text{CH}_3\text{COONa} \xrightarrow{\text{H}_2\text{O}} \text{CH}_3\text{COO}^- + \text{Na}^+$

-  $\text{CH}_3\text{COONa}$  fully dissociates in water to produce a high concentration of  $\text{CH}_3\text{COO}^-$  and  $\text{Na}^+$  ions, which are free to move and can carry charge, so  $\text{CH}_3\text{COONa}$  is a good conductor. //

## QUESTION TWO

Sufficient calcium carbonate,  $\text{CaCO}_3(\text{s})$ , is dissolved in water to make a saturated solution.

- (a) (i) Write the equation for the equilibrium occurring in a saturated solution of  $\text{CaCO}_3$ .



- (ii) Write the expression for  $K_s(\text{CaCO}_3)$ .

$$\text{// } K_s(\text{CaCO}_3) = [\text{Ca}^{2+}][\text{CO}_3^{2-}] \text{//}$$

- (iii) Calculate the solubility product of  $\text{CaCO}_3$ ,  $K_s(\text{CaCO}_3)$ .

The solubility of  $\text{CaCO}_3$  is  $5.74 \times 10^{-5} \text{ mol L}^{-1}$ .

$$\begin{aligned} \text{// } K_s(\text{CaCO}_3) &= (5.74 \times 10^{-5})^2 \\ &= 3.29 \times 10^{-9} \text{//} \end{aligned}$$

- (b) Some marine animals use calcium carbonate to form their shells. Increased acidification of the oceans poses a problem for the survival of these marine animals.

Explain why the solubility of  $\text{CaCO}_3$  is higher in an acidic solution.

Use an equation to support your explanation.

$\text{// CaCO}_3 \xrightarrow{\text{H}^+} \text{Ca}^{2+} + \text{CO}_3^{2-}$  -  $\text{CaCO}_3$  fully dissociates in water to form  $\text{Ca}^{2+}$  and  $\text{CO}_3^{2-}$  ions. In an acidic solution, there are  $\text{H}_3\text{O}^+$  ions present which will react with the  $\text{CO}_3^{2-}$  ions in solution to form a complex ion, therefore the forward reaction is favoured to produce more  $\text{CO}_3^{2-}$  to restore equilibrium, so the solubility of  $\text{CaCO}_3$  increases. //

~~CaCO3 + H+ -> Ca2+ + HCO3-~~

- (c) Show, by calculation, that a precipitate of lead(II) hydroxide,  $\text{Pb}(\text{OH})_2$ , will form when 25.0 mL of a sodium hydroxide solution,  $\text{NaOH}$ , at pH 12.6 is added to 25.0 mL of a 0.00421 mol  $\text{L}^{-1}$  lead(II) nitrate,  $\text{Pb}(\text{NO}_3)_2$ , solution.

$$K_s(\text{Pb}(\text{OH})_2) = 8.00 \times 10^{-17} \text{ at } 25^\circ\text{C}$$

$$\begin{aligned} // \text{ IP} &= [\text{A}^+][\text{B}^-] \\ &= [\text{NaOH}]^2 [\text{Pb}(\text{NO}_3)_2] \\ &= [\text{OH}^-]^2 [\text{Pb}^{2+}] \end{aligned}$$

$$\begin{aligned} 10^{-12.6} &= 2.51 \times 10^{-13} = [\text{H}_3\text{O}^+] \\ [\text{OH}^-] &= \frac{1 \times 10^{-14}}{2.51 \times 10^{-13}} \\ &= 0.0398 \text{ mol L}^{-1} \end{aligned}$$

∴

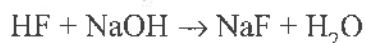
$$\begin{aligned} &= \left(0.0398 \times \frac{25}{50}\right)^2 \times \left(0.00421 \times \frac{25}{50}\right) \\ &= 8.34 \times 10^{-7} \text{ (3 s.f.)} \end{aligned}$$

IP >  $K_s$  so a precipitate will form //

## QUESTION THREE

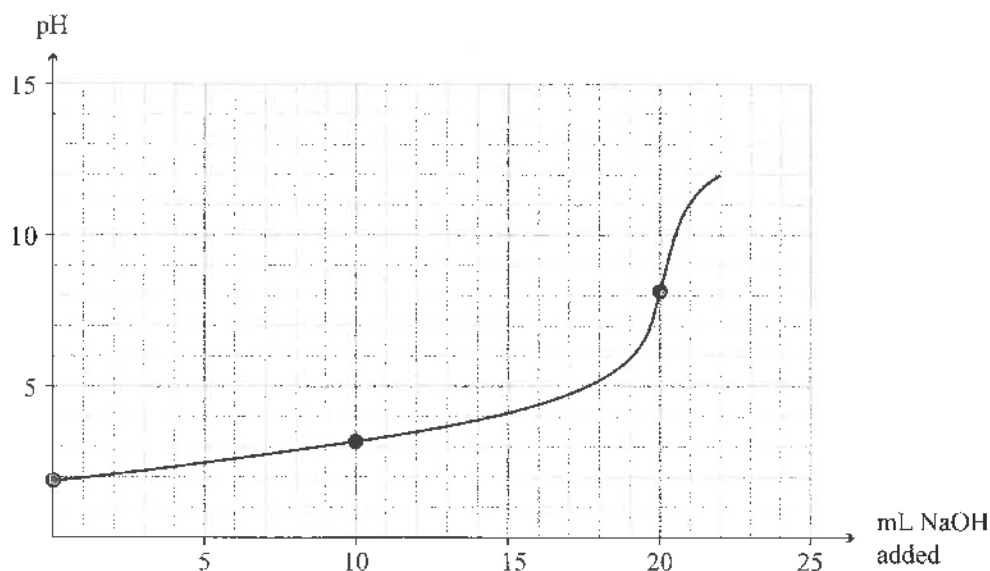
20.0 mL of  $0.258 \text{ mol L}^{-1}$  hydrofluoric acid, HF, solution is titrated with a sodium hydroxide, NaOH, solution.

The equation for the reaction is:



$$pK_a(\text{HF}) = 3.17$$

The titration curve is given below:



- (a) (i) Identify the species in solution at the equivalence point.

~~HF and H2O~~ //  $\text{Na}^+$ ,  $\text{F}^-$  and  $\text{H}_2\text{O}$  //

- (ii) Explain why the pH at the equivalence point is greater than 7.

Include an equation in your answer.

// At EQ pt, ~~the~~ the species present is ~~HF~~  $\text{F}^-$  which is the conjugate base of the weak acid, HF. This means the solution will be slightly basic (there are no more HF - all of it has reacted at EQ pt) so pH will be above 7.

$\text{F}^- + \text{H}_2\text{O} \rightleftharpoons \text{HF} + \text{OH}^-$  - the  $\text{F}^-$  in solution reacts with water to produce  $\text{OH}^-$  ions which increases pH. //

- (iii) After a certain volume of NaOH solution has been added, the concentration of HF in the solution will be twice that of the  $F^-$ .

Calculate the pH of this solution, and evaluate its ability to function as a buffer.

$$// \quad K_a = \frac{[H_3O^+] \times [F^-]}{[HF]}$$

$$K_a = \frac{[H_3O^+] \times [0.258 \times \frac{1}{2}]}{0.258}$$

$$K_a = 10^{-3.17} = 6.76 \times 10^{-4}$$

$$6.76 \times 10^{-4} = \frac{[H_3O^+] \times [0.258 \times \frac{1}{2}]}{0.258}$$

$$[H_3O^+] = 1.35 \times 10^{-3}$$

$$pH = -\log(1.35 \times 10^{-3}) \\ = 2.87 //$$

pH = 2.87 is within the buffer range of 2.17 - 4.17 so ~~the~~ the solution at this point will be able to function as a

buffer. It will resist a change in pH when an acid or base is added to solution.

- (iv) Determine by calculation, the pH of the solution after 24.0 mL of 0.258 mol L<sup>-1</sup> NaOH solution has been added.

$$// \quad [OH^-] = \frac{0.258 \times 4}{44}$$

$$= 0.0235 \text{ mol L}^{-1}$$

$$[H_3O^+] = \frac{1 \times 10^{-14}}{0.0235}$$

$$= 4.255 \times 10^{-13} \text{ mol L}^{-1}$$

$$pH = -\log(4.255 \times 10^{-13})$$

$$= 12.4 //$$

Question Three continues on the following page.

- (b) In a second titration, a  $0.258 \text{ mol L}^{-1}$  ethanoic acid,  $\text{CH}_3\text{COOH}$ , solution was titrated with the NaOH solution.

Contrast the expected pH at the equivalence point with the HF titration.

$$pK_a(\text{CH}_3\text{COOH}) = 4.76$$

No calculations are necessary.

$$// pK_a \text{ of } \text{CH}_3\text{COOH} = 4.76$$

$$pK_a \text{ of HF} = 3.17$$

$$K_a(\text{HF}) = 10^{-3.17} = 6.76 \times 10^{-4}$$

$$K_a(\text{CH}_3\text{COOH}) = 10^{-4.76} = 1.74 \times 10^{-5}$$

HF has a larger  $K_a$  value which means in equal concentrations of ~~their~~ their acids, (equal because both acids are  $0.258 \text{ mol L}^{-1}$ )

HF will have a larger  $[\text{H}_3\text{O}^+]$  due to a larger  $K_a$  value.

This means HF will have a smaller pH than ethanoic acid

because  $\text{pH} = -\log[\text{H}_3\text{O}^+]$  - at EQ pt, ethanoic acid will have a ~~higher~~ higher pH value than HF. //

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

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



Excellence exemplar for 91392 2015		Total score	21
Q	Grade score	Annotation	
1	E7	<p>This provides evidence for M5 because the equations and species list are correct in part (a)(i), (ii) &amp; (iii) and in (a) (iv) the calculation is correct. Also in (b), although they recognise pH depends on hydronium ion concentration and degree of dissociation for all three substances they do not compare the relative strengths of the weak bases. Had they done this they would have gained E8. For conductivity they relate conductivity the presence of ions and dissociation for all three substances and relate this to the concentrations of ions.</p>	
2	E7	<p>This provides evidence for M6 because they correctly write the solubility equation and expression in (a)(i) &amp; (ii) and calculate the solubility product in (a) (iii). In (b) the answer is confused stating <math>\text{CO}_3^{2-}</math> ions form a complex ion. Had they done this part correctly they would have gained an E8. In (c) they correctly calculate the answer.</p>	
3	E7	<p>This provides evidence for E7 because they correctly calculate the pH of the buffer but do not evaluate its ability and in (a)(iv) they correctly calculate the pH. Had they evaluated the buffer they may have gained excellence and E8 for the question.</p>	