

91392



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3

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## Level 3 Chemistry, 2017

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

2.00 p.m. Wednesday 15 November 2017

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

(a) Hydrogen fluoride, HF, and hydrogen bromide, HBr, both form acidic solutions when added to water.

(i) Write an equation for the reaction of each acid with water.

Hydrogen fluoride, HF, with water:



Hydrogen bromide, HBr, with water:



(ii) Compare and contrast the electrical conductivity of 0.150 mol L<sup>-1</sup> solutions of hydrofluoric acid, HF, and hydrobromic acid, HBr.

In your answer, you should:

- include the requirements for a solution to conduct electricity
- identify the species present AND their relative concentrations.

*No calculations are necessary.*

Electrical conductivity depends upon the amount of ions in solution that are able to carry charge. HF and HBr are both weak acids as they partially dissociate to form F<sup>-</sup> and H<sub>3</sub>O<sup>+</sup> in the case of HF and Br<sup>-</sup> and H<sub>3</sub>O<sup>+</sup> in the case of HBr. The equilibrium lies to the left. Each weak acid forms two ionic products which are able to carry charge, and in equal concentrations, so will have similar conductivity.

- (b) 40.0 mL of 0.150 mol L<sup>-1</sup> HBr solution was added to 25.0 mL of a saturated silver bromide, AgBr, solution.

- (i) Write an equation for the equilibrium occurring in a saturated solution of AgBr.



- (ii) Explain the changes that occur to the concentrations of the species in the saturated solution of AgBr on the addition of the HBr solution.

On addition, HBr dissociates to form H<sub>3</sub>O<sup>+</sup> and Br<sup>-</sup> ions (HBr + H<sub>2</sub>O ⇌ Br<sup>-</sup> + H<sub>3</sub>O<sup>+</sup>). According to Le Chatelier's principle, the excess of Br<sup>-</sup> ions will cause the AgBr equilibrium to favour the reverse reaction to reduce the concentration of Br<sup>-</sup> ions. Thus the concentration of Ag<sup>+</sup> and Br<sup>-</sup> decreases and concentration of AgBr increases. //

- (iii) Calculate the concentration of the silver ions, Ag<sup>+</sup>, after the HBr solution has been added.

$$K_s(\text{AgBr}) = 5.00 \times 10^{-13}$$

Assume the concentration of Br<sup>-</sup> in the original saturated solution of AgBr is insignificant.

$$[\text{Ag}^+][\text{Br}^-] = 5.00 \times 10^{-13}$$

$$c(\text{Br}^-) = \frac{n(\text{Br}^-)}{0.065}$$

$$[\text{Ag}^+][0.0923] = 5.00 \times 10^{-13}$$

$$n(\text{old Br}^-) = 0.150 \times 0.04$$

$$[\text{Ag}^+] = 5.42 \times 10^{-12} \text{ mol L}^{-1}$$

$$= 6 \times 10^{-3} \text{ mol}$$

$$c(\text{new Br}^-) = \frac{6 \times 10^{-3}}{0.065}$$

$$= 0.0923$$

## QUESTION TWO

(a) Ammonia,  $\text{NH}_3$ , is a weak base.

$$pK_a(\text{NH}_4^+) = 9.24$$

$$K_a(\text{NH}_4^+) = 5.75 \times 10^{-10}$$

(i) Calculate the pH of a  $0.105 \text{ mol L}^{-1}$   $\text{NH}_3$  solution.

$$\begin{aligned} [\text{H}_3\text{O}^+] &= \sqrt{\frac{K_a \times K_w}{C(\text{base})}} \\ &= \sqrt{\frac{5.75 \times 10^{-10} \times 10^{-14}}{0.105}} \\ &= 7.40 \times 10^{-12} \end{aligned}$$

$$\begin{aligned} -\log [7.40 \times 10^{-12}] &= \text{pH} \\ &= \underline{\underline{11.13}} // \end{aligned}$$

(ii) Dilute hydrochloric acid,  $\text{HCl}$ , is added to the  $\text{NH}_3$  solution until the ratio of  $\text{NH}_3$  to  $\text{NH}_4^+$  in the solution is 5:1.

Determine the pH of this solution, and evaluate its ability to resist a change in pH when small volumes of strong acid or base are added.

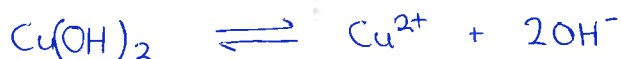
$$\begin{aligned} \text{pH} &= pK_a + \log \frac{[\text{base}]}{[\text{acid}]} \\ &= 9.24 + \log \frac{5}{1} \\ &= \underline{\underline{9.93}} \end{aligned}$$

The pH of the solution is 9.93 meaning there is a higher concentration of conjugate base  $\text{NH}_3$ , than  $\text{NH}_4^+$ . Thus, the solution will be more effective at maintaining pH when small volumes of strong acid are added as the excess  $\text{NH}_3$  reacts with it to form  $\text{NH}_4^+$  and water.





- (b) (i) Write the equation for the equilibrium occurring in a saturated solution of copper(II) hydroxide,  $\text{Cu}(\text{OH})_2$ .



- (ii) Write the expression for  $K_s(\text{Cu}(\text{OH})_2)$ .

$$K_s(\text{Cu}(\text{OH})_2) = [\text{Cu}^{2+}][\text{OH}^-]^2$$

- (iii) Calculate the solubility of  $\text{Cu}(\text{OH})_2$  in water at  $25^\circ\text{C}$ .

$$K_s(\text{Cu}(\text{OH})_2) = 4.80 \times 10^{-20}$$

$$[\text{Cu}^{2+}][\text{OH}^-]^2 = 4.80 \times 10^{-20}$$

$$4s^3 = 4.80 \times 10^{-20}$$

$$s = 2.29 \times 10^{-7} \text{ mol L}^{-1}$$

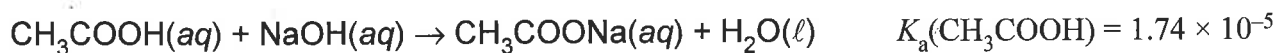
- (c) Explain why the solubility of  $\text{Cu}(\text{OH})_2$  increases when dilute hydrochloric acid is added.

$\text{Cu}(\text{OH})_2$  dissociates to form  $\text{Cu}^{2+}$  and  $\text{OH}^-$  ions  
 ( $\text{Cu}(\text{OH})_2 \rightleftharpoons \text{Cu}^{2+} + 2\text{OH}^-$ ). When dilute hydrochloric acid is added it forms  $\text{H}_3\text{O}^+$  and  $\text{Cl}^-$  ions ( $\text{HCl} + \text{H}_2\text{O} \rightleftharpoons \text{Cl}^- + \text{H}_3\text{O}^+$ ).  
 These  $\text{H}_3\text{O}^+$  ions neutralise  $\text{OH}^-$  ions to form  $\text{H}_2\text{O}$   
 ( $\text{OH}^- + \text{H}_3\text{O}^+ \rightleftharpoons \text{H}_2\text{O}$ ), which decreases the concentration of  $\text{OH}^-$  ions in the solution. Thus the  $\text{Cu}(\text{OH})_2$  equilibrium favours the forward reaction to replenish  $\text{OH}^-$  ions, increasing the solubility as more  $\text{Cu}(\text{OH})_2$  dissolves.

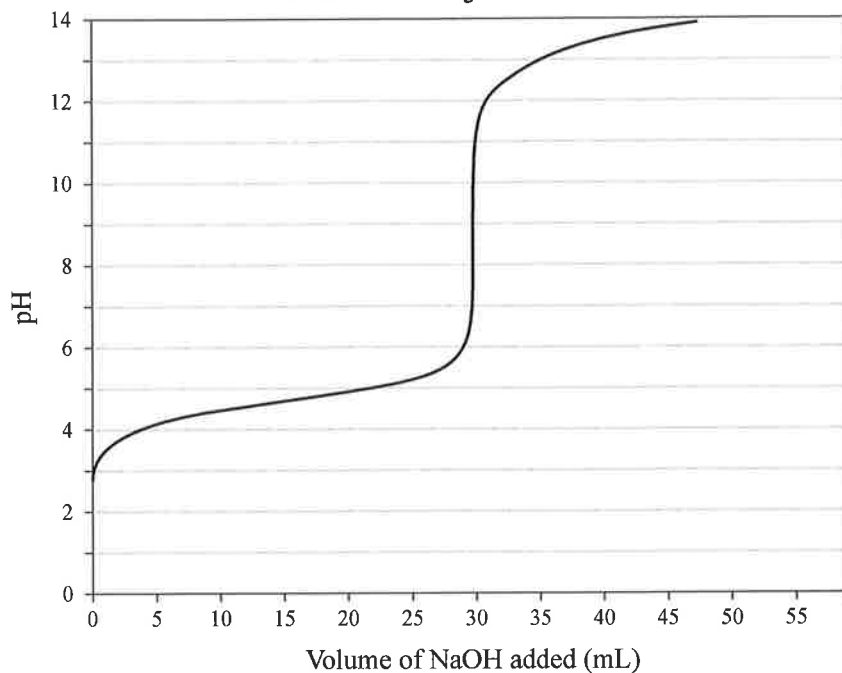
**QUESTION THREE**

A titration was carried out by adding  $0.112 \text{ mol L}^{-1}$  sodium hydroxide solution,  $\text{NaOH}(aq)$ , to  $20.0 \text{ mL}$  of ethanoic acid solution,  $\text{CH}_3\text{COOH}(aq)$ .

The equation for the reaction is:



**Titration curve for  $\text{CH}_3\text{COOH}$  versus  $\text{NaOH}$**



- (a) With reference to the titration curve above, put a tick next to the indicator most suited to identify the equivalence point.

Indicator	$pK_a$	Tick ONE box below
Methyl yellow	3.1	<input type="checkbox"/>
Bromocresol purple	6.3	<input checked="" type="checkbox"/>
Phenolphthalein	9.6	<input type="checkbox"/>

- (b) (i) The ethanoic acid solution,  $\text{CH}_3\text{COOH}(aq)$ , has a pH of 2.77 before any NaOH is added.

Show by calculation that the concentration of the  $\text{CH}_3\text{COOH}$  solution is  $0.166 \text{ mol L}^{-1}$ .



$$1.74 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+]^2}{[\text{CH}_3\text{COOH}]}$$

$$10^{-2.77} = [\text{H}_3\text{O}^+] = 1.698 \times 10^{-3}$$

$$[\text{CH}_3\text{COOH}] = \frac{(1.698 \times 10^{-3})^2}{1.74 \times 10^{-5}}$$

$$= 0.166 \text{ mol L}^{-1}$$

- (ii) Calculate the pH of the solution in the flask after 10.0 mL of  $0.112 \text{ mol L}^{-1}$  NaOH has been added to 20.0 mL of ethanoic acid solution,  $\text{CH}_3\text{COOH}(aq)$ .

$$\begin{aligned} n(\text{CH}_3\text{COOH}) &= \cancel{n(\text{CH}_3\text{COOH})} - n(\text{NaOH added}) \\ &= \cancel{(0.166 \times 0.02)} - (0.112 \times 0.01) \\ &= \cancel{0.00332} - 0.00112 = 2.2 \times 10^{-3} \end{aligned}$$

$$C(\text{CH}_3\text{COOH}) = \frac{2.2 \times 10^{-3}}{0.03} = 0.0733 \text{ mol L}^{-1}$$

$$C(\text{NaOH}) = \frac{n(\text{NaOH added})}{0.03} = \frac{(0.112 \times 0.01)}{0.03} = 0.0373 \text{ mol L}^{-1}$$

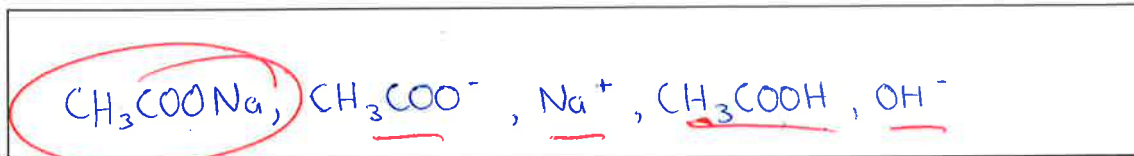
$$\begin{aligned} \text{pH} &= \text{pK}_a + \log \frac{[\text{base}]}{[\text{acid}]} \\ &= 4.76 + \log \frac{0.0373}{0.0733} \end{aligned}$$

$$= 4.47$$

Question Three continues on the following page.

(c) The equivalence point pH for the titration of ethanoic acid with sodium hydroxide is 8.79.

(i) Identify the chemical species present at the equivalence point, other than water.



(ii) In a second titration, a  $0.166 \text{ mol L}^{-1}$  methanoic acid solution,  $\text{HCOOH}(aq)$ , is titrated with the NaOH solution. The equivalence point pH for this titration is 8.28.

The equivalence point pH for the  $\text{CH}_3\text{COOH}$  titration is 8.79.

Compare and contrast the pH values at the equivalence point for both titrations.

$$K_a(\text{HCOOH}) = 1.82 \times 10^{-4} \quad K_a(\text{CH}_3\text{COOH}) = 1.74 \times 10^{-5}$$

No calculations are necessary.

HCOOH has a higher  $K_a$  value than  $\text{CH}_3\text{COOH}$ . This means that HCOOH will dissociate more than  $\text{CH}_3\text{COOH}$  i.e. the equilibrium lies further to the right ( $\text{HCOOH} + \text{H}_2\text{O} \rightleftharpoons \text{HCOO}^- + \text{H}_3\text{O}^+$ ). Thus there will be more  $\text{H}_3\text{O}^+$  ions in the solution, giving it a lower pH than  $\text{CH}_3\text{COOH}$ .



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

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QUESTION  
NUMBER

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

QUESTION  
NUMBER

ASSESSOR'S  
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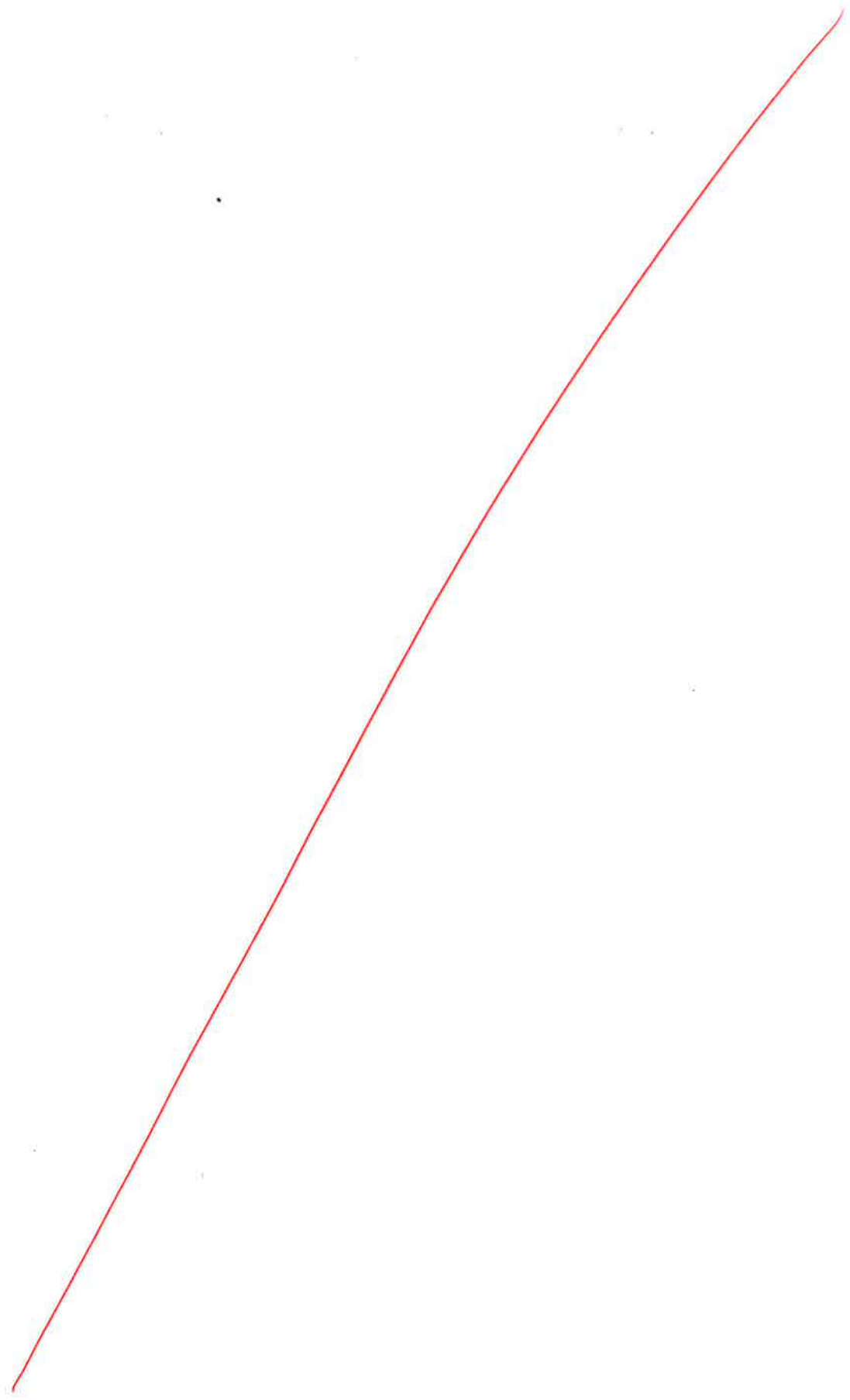
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91392





Excellence exemplar for 91392 2017		Total score	21
Q	Grade score	Annotation	
1	E7	<p>The candidate was awarded E7 for the following reasons:</p> <p>In part (a)(i), the correct arrow was used for hydrogen fluoride, but not for hydrogen bromide. In part (a)(ii), the candidate had the correct ideas on both conductivity and strength.</p> <p>In part (b), the candidate wrote the correct equation; had a full discussion on solubility – justifying their response using Le Chatelier’s Principle; correctly used <math>K_s</math> expression to correctly calculate the concentration of the silver ions present in the dilution.</p>	
2	E7	<p>The candidate was awarded E7 for the following reasons:</p> <p>In part (a), the candidate calculated the pH correctly, but the answer did not reflect the appropriate number of significant figures. The correct buffer pH and evaluation were given.</p> <p>In part (b), the correct equation and expression were given which were used to correctly calculate the solubility of <math>\text{Cu}(\text{OH})_2</math>, with the correct unit.</p> <p>In part (c), a full discussion on solubility relating to both equilibrium and acid-base reaction was given.</p>	
3	E7	<p>The candidate was awarded E7 for the following reasons:</p> <p>In part (a), an incorrect indicator was chosen.</p> <p>In part (b), the calculation of both concentration and pH were correct, with justification.</p> <p>In part (c), four correct species were given; the pH related to acid strength, however, to gain E8, the candidate’s response need to relate to all species present, e.g. the conjugate base strengths.</p>	