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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.

Merit

TOTAL 15

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QUESTION ONE

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(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:

$HF(l) + H_2O(l) \rightarrow H_3O^+(aq) + F^-(aq)$

Hydrogen bromide, HBr, with water:

$HBr(l) + H_2O(l) \rightarrow H_3O^+(aq) + Br^-(aq)$

(ii) Compare and contrast the electrical conductivity of 0.150 mol L^{-1} 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.

For solution to conduct electricity, there must be free moving charged electrons. HF and HBr are both strong acids which are fully dissolve into water. For example, 0.15 mol L^{-1} of HF and HBr solution will turn into 0.3 mol L^{-1} of charged ions after dissolve into water. In this case, HF solution will have H_3O^+ ions and F^- ions to conduct electrity. HBr solution will have Br^- ions and H_3O^+ ions to conduct electricity. Their electric ~~conducitity~~ conductivity will be similar as their concentration are same.

(b) 40.0 mL of 0.150 mol L^{-1} 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.

$AgBr(s) \rightleftharpoons Ag^{+}(aq) + Br^{-}(aq)$

(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.

Saturated solution of AgBr means no more Ag^+ ions or Br^- ions can dissolve into water. However, after adding HBr solution, the equilibrium will shift to left to minimise the change. Br^- ions' concentration will increase. Because no more Ag^+ ions can dissolve into water, $[Ag^+]$ will decrease.

(iii) Calculate the concentration of the silver ions, Ag^+, after the HBr solution has been added.

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

Assume the concentration of Br^- in the original saturated solution of AgBr is insignificant.

$k_s = [Ag^+(aq)][Br^-(aq)]$

Assume $[Ag^+] = X$, $[Br^-] = 0.15$

$k_s = 0.15X$

$X = 3.33 \times 10^{-12}$ mol L^{-1}

$[Ag^+] = 3.33 \times 10^{-12}$ mol L^{-1}

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QUESTION TWO

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(a) Ammonia, NH_3, is a weak base.

$pK_a(NH_4^+) = 9.24$ $K_a(NH_4^+) = 5.75 \times 10^{-10}$

(i) Calculate the pH of a 0.105 mol L^{-1} NH_3 solution.

$\cancel{NH_{3(aq)} + H_2O(l)}$ $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

$k_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{[OH^-]^2}{[NH_3]}$

$\cancel{Ka = 10^{-PKa} = 5.754 \times 10^{-10}}$

$k_b = \frac{kw}{ka} = \frac{1 \times 10^{-14}}{5.754 \times 10^{-10}} = 1.738 \times 10^{-5}$

Assume $[NH_3]$ at equilibrium $= [NH_3]$ at initial

$\therefore$ $1.738 \times 10^{-5} = \frac{[OH^-]^2}{0.105}$

$[OH^-] = 1.351 \times 10^{-3}$ mol L^{-1}

$[H_3O^+] = \frac{kw}{[OH^-]} = 7.4027 \times 10^{-12}$ mol L^{-1}

$PH = -\log[H_3O^+] = 11.1$

(ii) Dilute hydrochloric acid, HCl, is added to the NH_3 solution until the ratio of NH_3 to 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.

$\cancel{n(NH_3) : n(NH_4^+) = 5:1, n(NH_4^+) : n(OH^-) = 1:1}$

$NH_{3(aq)} + HCl_{(aq)} \rightleftharpoons NH_4^+{}_{(aq)} + Cl^-(aq)$

$C(NH_3) : C(NH_4^+) = 5:1$, $C(NH_4^+) : C(Cl^-) = 1:1$

$NH_4^+(aq) + H_2O \rightleftharpoons NH_3(aq) + H_3O^+(aq)$

$ka = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = [H_3O^+] \times 5$

$5.75 \times 10^{-10} = 5 \times [H_3O^+]$

$[H_3O^+] = 1.15 \times 10^{-10}$ mol L^{-1}

$PH = -\log[H_3O^+] = 9.94$

Because the solution is a mixture of strong acid and weak base. When small volume of strong acid or base are added, they will react with HCl or NH_3 and netralise to water. Hence, it won't affect PH value.

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

$Cu(OH)_2\ (s) \rightleftharpoons Cu^{2+}\ (aq) + 2OH^-\ (aq)$

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(ii) Write the expression for $K_s(Cu(OH)_2)$.

$K_s = [Cu^{2+}(aq)][OH^-(aq)]^2$

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(iii) Calculate the solubility of $Cu(OH)_2$ in water at 25°C.

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

Assume $[Cu^{2+}] = x$. $[OH^-] = 2x$

$K_s = 4x^3$

$4.8 \times 10^{-20} = 4x^3$

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

The solubility is 2.29×10^{-7} mol L^{-1}.

(c) Explain why the solubility of $Cu(OH)_2$ increases when dilute hydrochloric acid (HCl) is added.

$HCl\ (aq) + OH^-\ (aq) \rightarrow H_2O\ (aq) + Cl^-\ (aq)$

When dilute HCl is added, it will react with OH^- ions from $Cu(OH)_2$ to form water and Cl^- ions. Hence, the equilibrium reaction will favour right side to minimise the change. It means to increase the concentration of OH^- ions. In order to do that, the solubility of $Cu(OH)_2$ will increase.

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E7

QUESTION THREE

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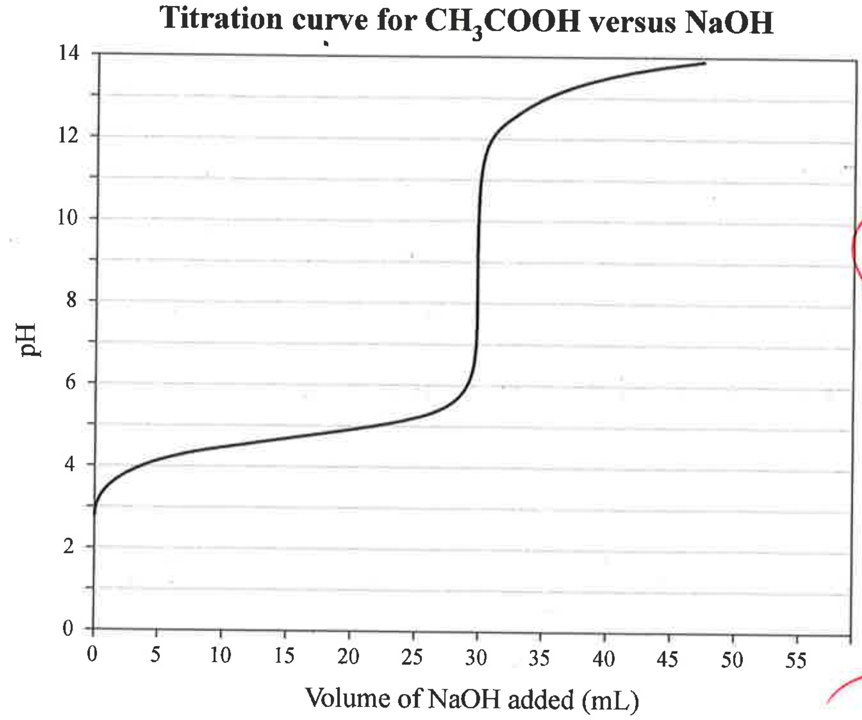
A titration was carried out by adding 0.112 mol L^{-1} sodium hydroxide solution, $NaOH(aq)$, to 20.0 mL of ethanoic acid solution, $CH_3COOH(aq)$.

strong base

The equation for the reaction is:

$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(\ell)$ $K_a(CH_3COOH) = 1.74 \times 10^{-5}$



(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	
Bromocresol purple	6.3	
Phenolphthalein	9.6	✓

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(b) (i) The ethanoic acid solution, $CH_3COOH(aq)$, has a pH of 2.77 before any NaOH is added.

Show by calculation that the concentration of the CH_3COOH solution is 0.166 mol L^{-1}.

$$CH_3COOH_{(aq)} + H_2O \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

$$ka = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} = \frac{[H_3O^+]^2}{[CH_3COOH]}$$

$$\therefore [CH_3COOH] = \frac{[H_3O^+]^2}{ka}$$

$$[H_3O^+] = 10^{-pH} = 10^{-2.77} = 1.698 \times 10^{-3} \text{ mol } L^{-1}$$

$$[CH_3COOH] = \frac{(1.698 \times 10^{-3})^2}{1.74 \times 10^{-5}} = 0.166 \text{ mol } L^{-1}$$

Assume $[CH_3COOH]$ at equilibrium = $[CH_3COOH]$ at initial

$\therefore$ $[CH_3COOH]$ at initial $= 0.166$ mol L^{-1} // 8

(ii) Calculate the pH of the solution in the flask after 10.0 mL of 0.112 mol L^{-1} NaOH has been added to 20.0 mL of ethanoic acid solution, $CH_3COOH(aq)$.

$$C_{(NaOH)} = 0.112 \times \frac{10}{20+10} = 0.0373 \text{ mol } L^{-1}$$

$$C_{(CH_3COOH)} = 0.166 \times \frac{20}{20+10} = 0.1101 \text{ mol } L^{-1}$$

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$$\cancel{CH_3COO^-_{(aq)} + NaOH_{(aq)} \rightleftharpoons CH_3COONa_{(aq)} + OH^-_{(aq)}}$$

$$\cancel{kb = \frac{[CH_3COONa][OH^-]}{[NaOH]}}$$

$$\cancel{CH_3COO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)}}$$

$$\cancel{kb = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]}}$$

$$CH_3COO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)}$$

$$kb = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]}$$

$$kb = \frac{kw}{ka} = \frac{1 \times 10^{-14}}{1.74 \times 10^{-5}} = 5.747 \times 10^{-10}$$

$$5.747 \times 10^{-10} = \frac{[OH^-] \times 0.1101}{0.0373}$$

$$[OH^-] = 1.70 \times 10^{-10} \text{ mol } L^{-1}$$

$$[H_3O^+] = 5.895 \times 10^{-6} \text{ mol } L^{-1}$$

$$pH = -\log [H_3O^+] = 5.23$$ //

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.

CH_3COO^- , Na^+ , OH^- , H_3O^+

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

The equivalence point pH for the CH_3COOH titration is 8.79.

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

$K_a(HCOOH) = 1.82 \times 10^{-4}$ > $K_a(CH_3COOH) = 1.74 \times 10^{-5}$

No calculations are necessary.

Equivalence point means right amount of NaOH is added to the HCOOH / ~~or~~ CH_3COOH solution so that all acid or base are netralised.

The pH value shows how much H_3O^+ / OH^- ions are present in the solution.

The pH value of CH_3COOH is slightly ~~to~~ higher than HCOOH means there're more OH^- ions present in the solution.

Because ka (HCOOH) > ka (CH_3COOH), therefore, PH (HCOOH) < PH (CH_3COOH).

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

QUESTION NUMBER

ASSESSOR'S USE ONLY

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

QUESTION NUMBER

ASSESSOR'S USE ONLY

Extra paper if required.
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QUESTION NUMBER

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Merit exemplar for 91392 2017		Total score	15
Q	**Grade score**	**Annotation**	
1	A4	The candidate was awarded A4 for the following reasons: In part (a)(i), the correct arrow was used for hydrogen bromide, but not for hydrogen fluoride. In part (a)(ii), the candidate had the correct ideas on both conductivity and strength. In part (b), the candidate wrote the correct equation; had a contradictory statement in the discussion on solubility; had a correct K_s expression, but incorrectly calculated the concentration of the silver ions present in the dilution.	
2	E7	The candidate was awarded E7 for the following reasons: In part (a), the candidate calculated the pH correctly, had the correct buffer pH, however, lacked the evaluation asked for in the question. In part (b), the equation had an incorrect arrow, however, the correct expression was given which was used to correctly calculate the solubility of $Cu(OH)_2$, with the correct unit. In part (c), a full discussion on solubility relating to both equilibrium and acid-base reaction was given.	
3	A4	The candidate was awarded A4 for the following reasons: In part (a), the correct indicator was chosen. In part (b), the calculation of concentration was justified, while one of the three steps in the calculation of the pH was correct. In part (c), the correct species were given; the relative pH's related to the hydroxide concentrations. To move up to merit, the candidate's response need to relate to all species present, e.g. the conjugate base strengths.	