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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.

Achievement

TOTAL 11

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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:

$HF + H_2O \rightarrow H_3O^+ + F^-$

Hydrogen bromide, HBr, with water:

$HBr + H_2O \rightleftharpoons H_3O^+ + Br^-$

(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.

- In order for a solution to conduct electricity, there must be ions present. A solution with a high concentration of ions will be a good conductor whereas a solution with a low concentration will be a poor conductor of electricity.
- Hydrofluoric acid is a very strong acid, which means it fully dissociates into its ions. This results in high concentrations of H_3O^+ and F^-, which therefore makes HF a good conductor of electricity.
- In contrast, hydrobromic acid is a weak acid, and only partially dissociates into its ions. This results in low concentrations of H_3O^+ and Br^-, which means fewer particles are present to carry a charge, and means that HBr is not a good conductor.

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(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.

$AgBr \rightleftharpoons Ag^+ + Br^-$

(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.

Due to the addition of HBr, more Bromine ions will be present in the solution. An increase in Br^- results in the equilibrium favouring the reactant (AgBr) side. And as the solution is saturated this will result in a precipitate of AgBr forming.

(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.

$K_s = [Ag^+][Br^-]$ 1:1 ratio

$K_s = s^2$

$s = \sqrt{5 \times 10^{-13}}$

$s = 7.071 \times 10^{-7}$

$n = cv$

$n = 7.071 \times 10^{-7} \times 0.025$

$n = 1.768 \times 10^{-8}$

$c = \frac{n}{v}$ $c = \frac{1.768 \times 10^{-8}}{0.065}$ $c = 2.72 \times 10^{-7}$

$\Rightarrow [Ag^+] = 2.72 \times 10^{-7}$

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pH = pKa + log (base / acid)

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.

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

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

Assuming that $[NH_4^+] = [OH^-]$

and that $[NH_3] = 0.105$ mol L^{-1}

$Ka = \frac{[OH^-]^2}{0.105}$ $[OH^-] = \sqrt{5.75 \times 10^{-10} \times 0.105}$

$[OH^-] = 7.77 \times 10^{-6}$ $[H_3O^+] = \frac{Kw}{[OH^-]}$

$[H_3O^+] = 1.29 \times 10^{-9}$

$-\log 1.29 \times 10^{-9} = 8.89 \Rightarrow pH = 8.89$

(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.

$[NH_3] = 5[NH_4^+]$

~~Assuming $[NH_3] = 0.105$ mol L^{-1}~~

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

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

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

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

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

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

$K_s = 4s^3$

$s = \sqrt[3]{\frac{4.8 \times 10^{-20}}{4}}$

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

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

When H_3O^+ ions are added they will react with the OH^- ions present to form water. As the product side of the equilibrium experiences a decrease, the product side will be favoured in order to oppose the change. This results in a decrease of $Cu(OH)_2$ as more is dissolved.

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A3

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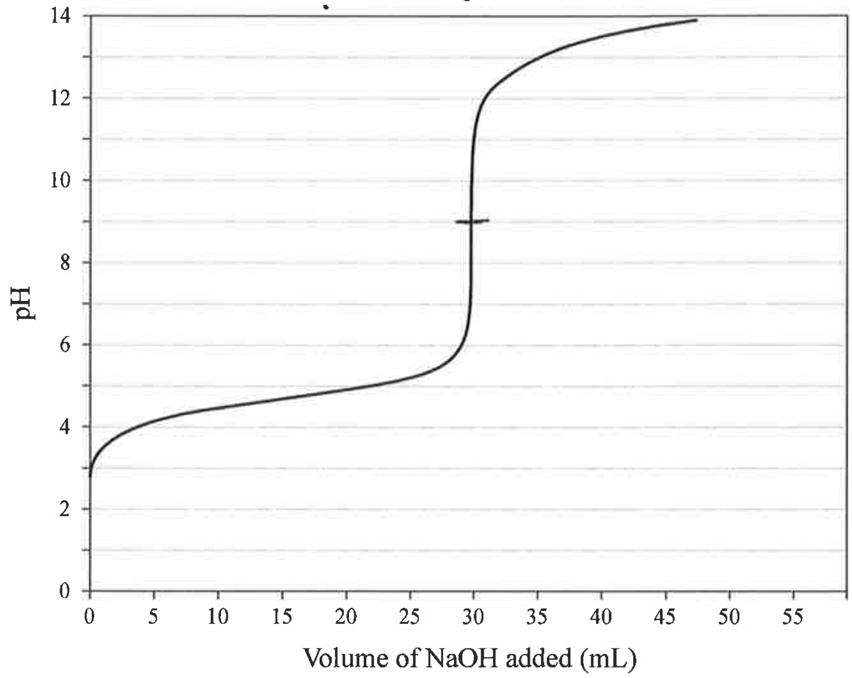
QUESTION THREE

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, $CH_3COOH(aq)$.

The equation for the reaction is:

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

Titration curve for CH_3COOH versus 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	
Bromocresol purple	6.3	
Phenolphthalein	9.6	✓

(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}.

$CH_3COOH \rightleftharpoons CH_3COO^- + H_3O^+$

$[H_3O^+] = 10^{-2.77} = 1.70 \times 10^{-3}$

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

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

Assume $[H_3O^+] = [CH_3COO^-]$

$[CH_3COOH] = \frac{(1.70 \times 10^{-3})^2}{1.74 \times 10^{-5}} = 0.166$ mol L^{-1}

(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, $CH_3COOH(aq)$.

$NaOH + CH_3COOH \longrightarrow CH_3COONa + H_2O$

$n = cv$ ~~$n = 0.01 \times 0.112 = 0.00112$ mol~~

~~$\frac{0.00112}{0.03} = 0.0373$ mol L^{-1}~~

~~$n = 0.166 \times 0.02$~~

~~$n = 3.32 \times 10^{-3}$ $\frac{3.32 \times 10^{-3}}{0.02} = 0.1??$ mol L^{-1}~~

$\longrightarrow n = 0.01 \times 0.112$ $n = 1.12 \times 10^{-3}$ mol

$\frac{1.12 \times 10^{-3}}{0.03} = 0.0373$ mol L^{-1}

1:1 ratio => $[OH^-] = 0.0373$ molL^{-1}

$[H_3O^+]$ for 20ml $= 1.70 \times 10^{-3}$

$n = 1.7 \times 10^{-3} \times 0.02 = 3.4 \times 10^{-5}$

$\frac{3.4 \times 10^{-5}}{0.03} = 1.13 \times 10^{-3}$ molL^{-1}

$0.0373 \times 1.13 \times 10^{-3} = 4.23 \times 10^{-5}$

$-\log 4.23 \times 10^{-5} =$ ~~4.368~~ 4.37

Question Three continues on the following page.

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(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.

EOH$^-$ CH_3COONa NaOH

(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.

methanoic acid has a lower equivalence point pH as it dissociates better into its ions then ethanoic acid. This can be seen by the Ka values as HCOOH has a higher Ka then CH_3COOH, a higher Ka means more ions are present relative to the concentration of the solution. This means HCOOH has a lower pH then CH_3COOH, and results in the equivalence point being lower as well. As a stronger acid with the same base results in the equivalence point pH shifting more to 7.

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A4

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

QUESTION NUMBER

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

QUESTION NUMBER

ASSESSOR'S USE ONLY

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QUESTION NUMBER

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Achievement exemplar for 91392 2017		Total score	11
Q	**Grade score**	**Annotation**	
1	A4	The candidate was awarded A4 for the following reasons: In part (a)(i), both equations had incorrect arrows. 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	A3	The candidate was awarded A3 for the following reasons: In part (a), the candidate calculated the pH incorrectly and did not answer part (a)(ii). Unfortunately, this response was required for the candidate to move up into Merit or higher. In part (b), the correct equation and expression were given which were 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 relative pH's related to acid strength. To move up to merit, the candidate's response need to relate to all species present, e.g. the conjugate base strengths.	