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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

Methylammonium chloride, CH_3NH_3Cl, dissolves in water to form a weakly acidic solution.

$K_a(CH_3NH_3^+) = 2.29 \times 10^{-11}$

(a) (i) Write an equation to show CH_3NH_3Cl dissolving in water.

// $CH_3NH_3Cl \xrightarrow{H_2O} CH_3NH_3^+{}_{(aq)} + Cl^-_{(aq)}$ //

(ii) Write an equation to show the reaction occurring in an aqueous solution of CH_3NH_3Cl.

// $CH_3NH_3^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3NH_{2(aq)} + H_3O^+_{(aq)}$ //

(iii) List all the species present in an aqueous solution of CH_3NH_3Cl, in order of decreasing concentration.

Do not include water.

~~$CH_3NH_3^+ \approx Cl^- > CH_3NH_2 \approx H_3O^+$~~

~~$CH_3NH_3^+, Cl^-, CH_3NH_2, H_3O^+$~~ // $CH_3NH_3^+ > CH_3NH_2 \approx H_3O^+ > OH^-$ //

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(iv) Calculate the pH of 0.0152 mol L^{-1} CH_3NH_3Cl solution.

~~Ka [H3O+]~~ // $Ka = \dfrac{[H_3O^+]^2}{[CH_3NH_3Cl]}$

$2.29 \times 10^{-11} = \dfrac{[H_3O^+]^2}{0.0152}$

$[H_3O^+] = 5.90 \times 10^{-7}$ molL^{-1} (3s.f.)

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

$= 6.23$ (3s.f.) //

i

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

Solution	NaOH	CH_3NH_2	CH_3COONa
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: // $NaOH \xrightarrow{H_2O} Na^+ + OH^-$ - NaOH fully dissociates in water because it is a strong base, to produce a high concentration of Na^+ and OH^- ions. Because the concentration of OH^- ions produced is large, it has a high pH.

$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$ - CH_3NH_2 is a weak base so only partially dissociates in water. The reaction is reversible therefore there is only a small concentration of OH^- ions so it is only slightly basic (smaller pH than NaOH).

$CH_3COONa \xrightarrow{H_2O} CH_3COO^- + Na^+$, $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$

- CH_3COONa fully dissociates in water to produce CH_3COO^- and Na^+. CH_3COO^- reacts with water to form CH_3COOH and OH^-, however the reaction is reversible because CH_3COO^- is a weak base, so there are only a small concentration of OH^- ions in solution, so the pH is lower than NaOH. //

i

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

Electrical conductivity: * $NaOH \xrightarrow{H_2O} Na^+ + OH^-$ - NaOH fully dissociates in water to produce a high concentration of Na^+ and OH^- ions which are free to move and can carry charge, so NaOH is a good conductor of electricity. $CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$ - CH_3NH_2 only partially dissociates in water to produce a low concentration of $CH_3NH_3^+$ and OH^- ions, so there are a smaller concentration of ions that are free to move and can carry charge, so CH_3NH_2 is a poor conductor. $CH_3COONa \xrightarrow{H_2O} CH_3COO^- + Na^+$ - CH_3COONa fully dissociates in water to produce a high concentration of CH_3COO^- and Na^+ ions, which are free to move and can carry charge, so CH_3COONa is a good conductor. //

C

E7

QUESTION TWO

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Sufficient calcium carbonate, $CaCO_3(s)$, is dissolved in water to make a saturated solution.

(a) (i) Write the equation for the equilibrium occurring in a saturated solution of $CaCO_3$.

//$CaCO_{3(s)} \rightleftharpoons Ca^{2+}_{(aq)} + CO^{2-}_{3(aq)}$//

(ii) Write the expression for $K_s(CaCO_3)$.

//$K_s(CaCO_3) = [Ca^{2+}][CO_3^{2-}]$//

u

(iii) Calculate the solubility product of $CaCO_3$, $K_s(CaCO_3)$.

The solubility of $CaCO_3$ is 5.74×10^{-5} mol L^{-1}.

//$K_s(CaCO_3) = (5.74 \times 10^{-5})^2$
$= 3.29 \times 10^{-9}$//

i

(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 $CaCO_3$ is higher in an acidic solution.

Use an equation to support your explanation.

//$CaCO_3 \xrightarrow{H_2O} Ca^{2+} + CO_3^{2-}$ - $CaCO_3$ fully dissociates in water to form Ca^{2+} and CO_3^{2-} ions. In an acidic solution, there are H_3O^+ ions present which will react with the CO_3^{2-} ions in solution to form a complex ion, therefore the forward reaction is favoured to produce more CO_3^{2-} to restore equilibrium, so the solubility of $CaCO_3$ increases.//

(c) Show, by calculation, that a precipitate of lead(II) hydroxide, $Pb(OH)_2$, will form when 25.0 mL of a sodium hydroxide solution, NaOH, at pH 12.6 is added to 25.0 mL of a 0.00421 mol L^{-1} lead(II) nitrate, $Pb(NO_3)_2$, solution.

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

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IP = [A⁺][B⁻]

$= \cancel{[NaOH][Pb(NO_3)_2]}$

$= [OH^-]^2 [Pb^{2+}]$

$10^{-12.6} = 2.51 \times 10^{-13} = [H_3O^+]$

$[OH^-] = \frac{1 \times 10^{-14}}{2.51 \times 10^{-13}}$

$= 0.0398$ molL^{-1}

$= (0.0398 \times \frac{25}{50})^2 \times (0.00421 \times \frac{25}{50})$

$= 8.34 \times 10^{-7}$ (3s.f.)

IP > Ks so a precipitate will form.

C

E7

QUESTION THREE

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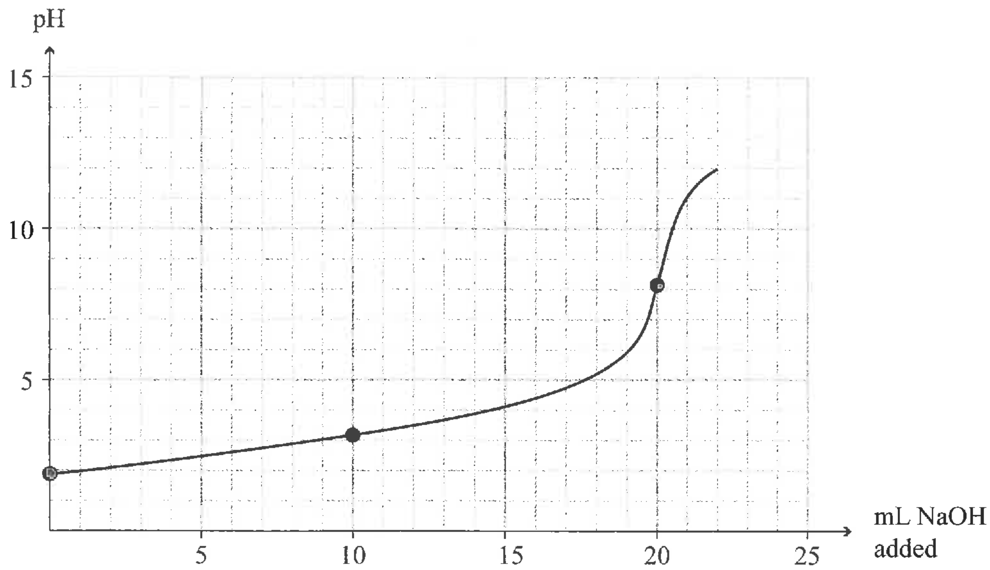
20.0 mL of 0.258 mol L^{-1} hydrofluoric acid, HF, solution is titrated with a sodium hydroxide, NaOH, solution.

The equation for the reaction is:

$$HF + NaOH \rightarrow NaF + H_2O$$

$$pK_a(HF) = 3.17$$

The titration curve is given below:



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

~~NaF and H₂O~~ // Na^+, F^- and H_2O //

u

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

Include an equation in your answer.

// At EQ pt, the species present is 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.

$F^- + H_2O \rightleftharpoons HF + OH^-$ – the F^- in solution reacts with water to produce OH^- ions which increases pH. //

i

(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.

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

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

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

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

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

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

$= 2.87$ //

pH = 2.87 is within the buffer range of 2.17 – 4.17 so the 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.

i

(iv) Determine by calculation, the pH of the solution after 24.0 mL of 0.258 mol L^{-1} NaOH solution has been added.

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

$= 0.0235$ molL^{-1}

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

$= 4.255 \times 10^{-13}$ molL^{-1}

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

$= 12.4$ //

C

Question Three continues on the following page.

(b) In a second titration, a 0.258 mol L^{-1} ethanoic acid, CH_3COOH, solution was titrated with the NaOH solution.

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

$pK_a(CH_3COOH) = 4.76$

No calculations are necessary.

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// pKa of $CH_3COOH = 4.76$

pKa of HF = 3.17

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

$Ka(CH_3COOH) = 10^{-4.76} = 1.74 \times 10^{-5}$

HF has a larger Ka value which means in equal concentrations of ~~~~ their acids, (equal because both acids are 0.258 molL^{-1}) HF will have a larger $[H_3O^+]$ due to a larger Ka value. This means HF will have a smaller pH than ethanoic acid because $pH = -\log[H_3O^+]$ – at EQ pt, ethanoic acid will have a ~~~~ higher pH value than HF.//

$Ka\uparrow = \frac{[H_3O^+]^2\uparrow}{[HF]}$

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

u

E7

Excellence exemplar for 91392 2015			Total score	21
Q	**Grade score**	**Annotation**		
1	E7	This provides evidence for M5 because the equations and species list are correct in part (a)(i), (ii) & (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.		
2	E7	This provides evidence for M6 because they correctly write the solubility equation and expression in (a)(i) & (ii) and calculate the solubility product in (a) (iii). In (b) the answer is confused stating CO_3^{2-} ions form a complex ion. Had they done this part correctly they would have gained an E8. In (c) they correctly calculate the answer.		
3	E7	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.		