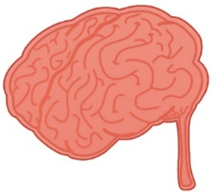
ANSWERS



Level 3 Chemistry

91392 Demonstrate understanding of equilibrium principles in aqueous systems

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

You should attempt ALL the questions in this booklet.

A periodic table is provided in the Resource Sheet.

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 – 9 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.

TOTAL

ASSESSOR'S USE ONLY

QUESTION ONE

8.00×10^{-3} g of calcium fluoride, CaF_2, will dissolve in 500 mL of water.

$M(CaF_2) = 78.0$ g mol^{-1}.

(a) Write the solubility product expression, K_s, for calcium fluoride.

$K_s = [Ca^{2+}][F^-]^2$

(b) (i) Calculate the solubility of calcium fluoride in mol L^{-1}, at this temperature.

8.00×10^{-3} g of CaF_2 will dissolve in 500 mL

0.0160 g of CaF_2 will dissolve in 1000 mL

$n = m/M$ $n = 0.0160/78.0 = 2.05 \times 10^{-4}$ mol L^{-1} (3 s.f.)

(ii) Calculate the K_s of calcium fluoride.

$Ks = 4s^3$

$Ks = 4s^3 = 3.45 \times 10^{-11}$ (3 s.f.)

(c) Explain how the solubility of calcium fluoride, CaF_2, will change if added to 500 mL of a 0.200 mol L^{-1} sodium fluoride solution instead of 500 mL of water.
Support your answer with balanced equations.
No calculations are necessary.

It will decrease as NaF contains fluoride, $F^-(aq)$ which is a common ion.
Using equilibrium principles, the back reaction will be favoured so the solubility of the CaF_2 will decrease.

(d) Determine whether a precipitate will form when 20.0 mL of 0.00200 mol L^{-1} potassium chromate, K_2CrO_4 are mixed with 60.0 mL of 1.25×10^{-4} mol L^{-1} lead nitrate, $Pb(NO_3)_2$. $K_s(PbCrO_4) = 2.8 \times 10^{-13}$ at 25°C.

IP or Q = $[Pb^{2+}(aq)][CrO_4^{2-}(aq)]$

$[Pb^{2+}(aq)]$ = 20.0/80.0 x 0.00200 = 5.00×10^{-4} mol L^{-1}

$[CrO_4^{2-}(aq)]$= 60.0/80.0 x 1.25×10^{-4} = 9.375×10^{-5} mol L^{-1}

Q = $[5.00 \times 10^{-4}][9.375 \times 10^{-5}] = 4.69 \times 10^{-8}$ (3 s.f.)

Since Q > K_s (2.8×10^{-13}) then a precipitate will form

Question Two

(a) Methyl ammonium bromide, CH_3NH_3Br dissolves in water to form a weakly acidic solution. $K_a(CH_3NH_3^+)$ is 2.28×10^{-11}.

Write an equation for when methyl ammonium bromide dissolves in water AND the reaction that then occurs in the aqueous solution.

$CH_3NH_3Br(s) \rightarrow CH_3NH_3^+(aq) + Br^-(aq)$

$CH_3NH_3^+(aq) + H_2O(l) \rightleftharpoons CH_3NH_2(aq) + H_3O^+(aq)$

Since $[H_3O^+(aq)] > [OH^-]$ from the dissociation of water then the solution is weakly acidic.

(b) Calculate the pH of 0.350 mol L^{-1} CH_3NH_3Br solution.

$K_a = [H_3O^+]^2 / [salt]$

$2.28 \times 10^{-11} = [H_3O^+]^2 / [0.350]$ therefore $[H_3O^+]^2 = 7.98 \times 10^{-12}$

$[H_3O^+] = 2.82 \times 10^{-6}$ mol L^{-1}. pH = $-\log 2.82 \times 10^{-6}$ = 5.55 (3 s.f.)

Or $K_a = [CH_3NH_2][H_3O^+] / [CH_3NH_3^+]$

(c) The table shows the pH and electrical conductivity of three solutions. Their concentrations are all 0.150 mol L^{-1}.

Solution	$C_6H_5NH_2$	NaOH	NH_4Cl
pH	8.91	13.2	5.06
Electrical conductivity	poor	good	good

Compare and contrast the pH and electrical conductivity of these three solutions. Include appropriate equations in your answer.

pH:

The pH of a solution is calculated from its $[H_3O^+]$.
NaOH is an ionic solid that is a strong base and dissociates completely to produce a high OH^- concentration (low $[H_3O^+]$).
Since $[OH^-]$ is high / $[H_3O^+]$ is low, the pH is high.
$NaOH \rightarrow Na^+ + OH^-$

$C_6H_5NH_2$ is a weak base that partially reacts / dissociates / ionises with H_2O producing a lower concentration of OH^-
Therefore it has a lower pH than NaOH:
$C_6H_5NH_2 + H_2O \rightleftharpoons C_6H_5NH_3^+ + OH^-$

The NH_4Cl is an ionic solid that dissolves completely in H_2O. The NH_4^+ ion is a weak acid that partially reacts with H_2O producing a low concentration of H_3O^+
$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$
The pH is closer to 7, showing it is a weak acid. Therefore it has a lowest pH.

Electrical conductivity:

Electrical conductivity is determined by the total concentration of ions.

Since $C_6H_5NH_2$ is a weak base, it only partially reacts with water to produce a low concentration of ions in solution so it is a poor electrical conductor.
$C_6H_5NH_2 + H_2O \rightleftharpoons C_6H_5NH_3^+ + OH^-$

NaOH completely dissolves to produce a high concentration of Na^+ and OH^- ions in solution. $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$ Therefore it is a good conductor.

NH4Cl is also an ionic solid. It dissolves completely to produce a high concentration of NH_4^+ and Cl^- ions: $NH_4Cl \rightarrow NH_4^+ + Cl^-$ Therefore it is a good conductor.

Question Three

Methanoic acid, HCOOH, is a weak acid. $pK_a(HCOOH) = 3.75$

(a) (i) List all the species present in a solution of HCOOH, in order of decreasing concentration. Do not include water.

$HCOOH > HCOO^- = H_3O^+ > OH^-$

(ii) Show that the pH of a 0.100 mol L^{-1} methanoic acid solution is 2.38.

$K_a = [H_3O^+]^2 / [acid]$

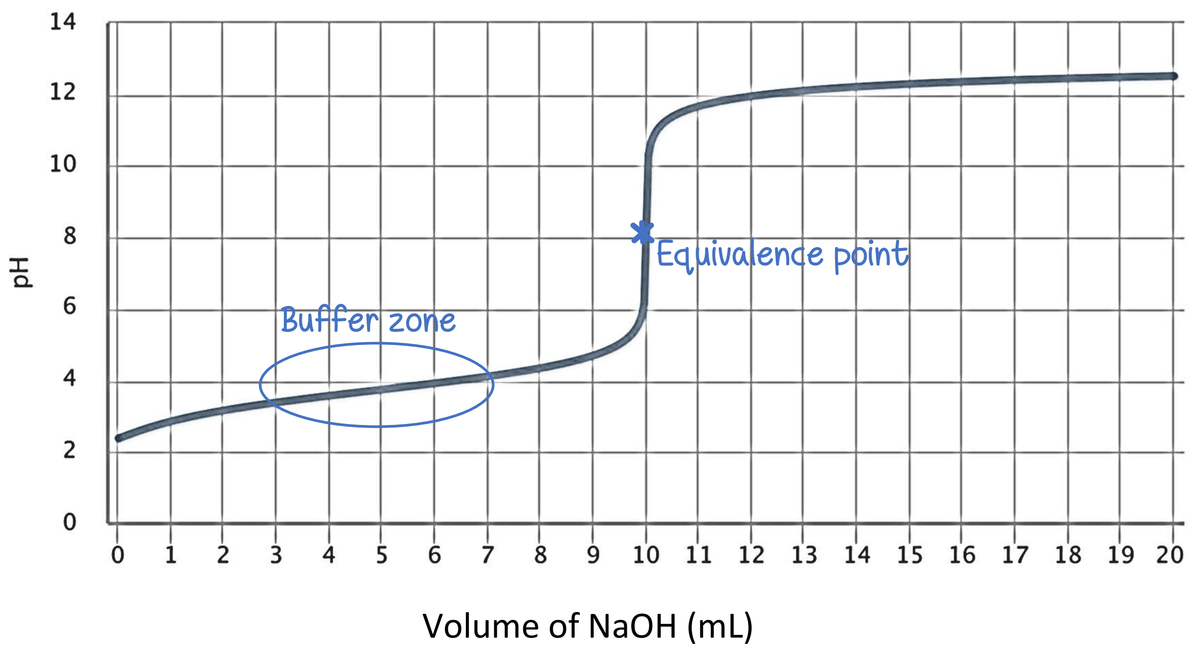
$pK_a(HCOOH) = 3.75$ therefore $K_a = 10^{-3.75} = 1.78 \times 10^{-4}$

$1.78 \times 10^{-4} = [H_3O^+]^2 / [0.100]$

$[H_3O^+]^2 = 1.78 \times 10^{-5}$, so $[H_3O^+] = 4.22 \times 10^{-3}$

$pH = -\log 4.22 \times 10^{-3} = 2.38$ (3 s.f.)

(b) (i) Here is the titration curve for the addition of 20.0 mL of 0.100 mol L^{-1} sodium hydroxide solution to 10.0 mL of 0.100 mol L^{-1} methanoic acid solution. Clearly label the buffer zone and equivalence point.



(ii) Identify the indicator that would be most suitable for this titration and explain your choice.

Indicator	Bromocresol green	Phenolphthalein	Alizarin yellow R
pK_a	4.7	9.4	11.2

Indicators are effective in the range $pH = pK_a \pm 1$.
Phenolphthalein is a suitable indicator as its pK_a is within 1 pH unit of equivalence point. Hence it will change colour at the equivalence point of the reaction in the steepest part of the graph.

Note: Bromocresol green will change colour in the buffer region and just after as its colour changes between pH 3.7 and 5.7 making this indicator unsuitable. Alizarin yellow R would change colours over pH range 10.2 -12.2 which is after the vertical portion.

(c) A buffer solution is formed when sodium hydroxide solution is added to methanoic acid. Using equations involving methanoate ions, describe how a solution containing methanoic acid and sodium methanoate acts as a buffer.

A buffer is a solution that undergoes a minimal change of pH when small amounts of acid or base are added.

HCOOH will react with added OH^- ions so there is almost no change in $[OH^-]$

$HCOOH + OH^- \rightarrow HCOO^- + H_2O$

$HCOO^-$ will react with added H_3O^+ ions so there is almost no change in $[H_3O^+]$

$HCOO^- + H_3O^+ \rightarrow HCOOH + H_2O$

(d) (i) 2.45 g of sodium hydroxide was added to 400 mL of 0.350 mol L^{-1} methanoic acid. Calculate the pH of the buffer. $M(NaOH) = 40.0$ g mol^{-1}.

pH = pKa + log ([base]/[acid])

$pK_a(HCOOH) = 3.75$

$NaOH + HCOOH \rightarrow HCOONa + H_2O$

n(NaOH) = m/M = 2.45 / 40.0 = 0.06125 mol.

n(HCOOH) = CV = 0.350 x 0.400 = 0.140 mol

n(HCOONa) made = 0.06125 mol

n(HCOOH) unreacted = 0.140 – 0.06125 = 0.07875 mol

c(HCOONa) 0.06125 / 0.400 = 0.153 mol L^{-1}*
c(HCOOH) 0.07875 / 0.400 = 0.197 mol L^{-1}*
*No need to calculate as it is ratio of mol base/acid that is needed and both now in same 400 mL)

pH = pka + log [base]/[acid] = 3.75 + log (0.153/0.197) = 3.64 (3 s.f.)

(ii) Evaluate the ability of this solution, in (d)(i) to function as a buffer.

It can work as a buffer as it contains both HCOOH and $HCOO^-$. The lower amount / concentration / number of $HCOO^-$ ions in solution mean it would not be as effective at buffering added H_3O^+ as it would be for added OH^- ions.

Extra paper if required.

Write the question number(s) if applicable

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