ACIDS PRACTICE pK_a and pK_b

1 Methanoic acid (HCOOH) is a weak acid with pK_a=3.75.

(a) Calculate the pH of a 0.100 mol dm^{-3} solution of methanoic acid. [3]

$K_a = 10^{-3.75} = 1.78\times10^{-4}$

$HCOOH \rightleftharpoons HCOO^- + H^+$

$K_a = \frac{[HCOO^-][H^+]}{[HCOOH]} = \frac{[H^+]^2}{0.100}$ assuming dissociation of acid negligible compared to its conc.

$[H^+]^2 = 0.100 \times 1.78\times10^{-4}$ $[H^+] = 4.22\times10^{-3}$ mol dm^{-3}

$pH = -\log_{10}[H^+] = 2.38$

(b) Deduce the pK_b value of the methanoate ion [1]

$pK_w = pK_a + pK_b$ 14 – 3.75 = 10.25 $pK_b = 10.25$

(c) Calculate the pH of a 0.100 mol dm^{-3} solution of sodium methanoate. [3]

$K_b = 10^{-10.25} = 5.62\times10^{-11}$

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

$K_b = \frac{[HCOOH][OH^-]}{[HCOO^-]} = \frac{[OH^-]^2}{0.100}$ assuming ionisation of base negligible compared to its conc.

$[OH^-]^2 = 0.100 \times 5.62\times10^{-11}$ $[OH^-] = 2.37\times10^{-6}$ mol dm^{-3}

$pOH = -\log_{10}[OH^-] = 5.625$ $pH = 14 - pOH = 14 - 5.625 = 8.375$

(d) Write an equation for the reaction between sodium hydroxide and methanoic acid. [1]

$HCOOH + NaOH \rightarrow HCOONa + H_2O$

(e) 25.00 cm^3 of a 0.100 mol dm^{-3} solution of methanoic acid is reacted with 25.00 cm^3 of a 0.100 mol dm^{-3} solution of sodium hydroxide. Deduce the concentration of sodium methanoate in the resulting solution. [1]

0.100/2 = 0.0500 mol dm^{-3}

same number of moles of sodium methanoate as of methanoic acid or sodium hydroxide but present in twice the volume of solution

(f) A solution of methanoic acid is titrated with sodium hydroxide solution. 25.00 cm^3 of the methanoic acid solution reacted exactly with 27.40 cm^3 of 0.100 mol dm^{-3} NaOH(aq). Calculate the pH of the resulting solution. [4]

Moles NaOH = 27.40/1000 x 0.100 = 2.740×10^{-3} mol - same as number of moles of methanoic acid and sodium methanoate formed.

Total volume = 25.00 + 27.40

Concentration of sodium methanoate = 2.740×10^{-3}/(52.40/1000) = 0.0523 mol dm^{-3}

$K_b = 10^{-10.25} = 5.62\times10^{-11}$ $HCOO^- + H_2O \rightleftharpoons HCOOH + OH^-$

$K_b = \frac{[HCOOH][OH^-]}{[HCOO^-]} = \frac{[OH^-]^2}{0.0523}$ assuming ionisation of base negligible compared to its conc.

$[OH^-]^2 = 0.0523 \times 5.62\times10^{-11}$ $[OH^-] = 1.71\times10^{-6}$ mol dm^{-3}

$pOH = -\log_{10}[OH^-] = 5.766$ $pH = 14 - pOH = 14 - 5.625 = 8.23$

2 Methylamine (CH_3NH_2) is a weak base with pK_b=3.36.

(a) Calculate the pH of a 0.0100 mol dm^{-3} solution of methylamine [3]

$K_b = 10^{-3.36} = 4.365\times10^{-4}$

$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$

$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]} = \frac{[OH^-]^2}{0.0100}$ assuming ionisation of base negligible compared to its conc.

$[OH^-]^2 = 0.0100 \times 4.365\times10^{-4}$ $[OH^-] = 2.089\times10^{-3}$ mol dm^{-3}

$pOH = -\log_{10}[OH^-] = 2.68$ $pH = 14 - pOH = 14 – 2.68 = 11.32$

(b) Deduce the pK_a value of the methylammonium ion ($CH_3NH_3^+$) [1]

$pK_w = pK_a + pK_b$ $14 – 3.36 = 10.64$ $pK_a = 10.64$

(c) Calculate the pH of a 0.0100 mol dm^{-3} solution of methylammonium chloride (CH_3NH_3Cl) [3]

$K_a = 10^{-10.64} = 2.29\times10^{-11}$

$CH_3NH_3^+ \rightleftharpoons CH_3NH_2 + H^+$

$K_a = \frac{[CH_3NH_2][H^+]}{[CH_3NH_3^+]} = \frac{[H^+]^2}{0.0100}$ assuming dissociation of acid negligible compared to its conc.

$[H^+]^2 = 0.0100 \times 2.29\times10^{-11}$ $[H^+] = 4.786\times10^{-7}$ mol dm^{-3}

$pH = -\log_{10}[H^+] = 6.32$

(d) Write an equation for the reaction between hydrochloric acid and methylamine. [1]

$CH_3NH_2 + HCl \rightarrow CH_3NH_3Cl$

(e) A solution of methylamine is titrated with hydrochloric acid. 25.00 cm^3 of the methylamine solution reacted exactly with 23.70 cm^3 of 0.0100 mol dm^{-3} HCl(aq). Calculate the pH of the resulting solution. [5]

Moles HCl = 23.70/1000 x 0.0100 = 2.370×10^{-4} mol - same as number of moles of methylamine and methylammonium chloride.

Concentration of methylammonium chloride = 2.370×10^{-4}/(48.70/1000) = 4.867×10^{-3} mol dm^{-3}

Total volume = 25.00 + 23.70

$K_a = 10^{-10.64} = 2.29\times10^{-11}$

$CH_3NH_3^+ \rightleftharpoons CH_3NH_2 + H^+$

$K_a = \frac{[CH_3NH_2][H^+]}{[CH_3NH_3^+]} = \frac{[H^+]^2}{4.867\times10^{-3}}$ assuming dissociation of acid negligible compared to its conc.

$[H^+]^2 = 4.867\times10^{-3} \times 2.29\times10^{-11}$ $[H^+] = 3.338\times10^{-7}$ mol dm^{-3}

$pH = -\log_{10}[H^+] = 6.48$