AS91393 Demonstrate understanding of equilibrium principles in aqueous systems
pH calculation questions from past papers

(2018:1)

(a) When sodium ethanoate, CH_3COONa, is dissolved in water, the resulting solution has a pH greater than 7 due to the production of hydroxide ions, OH^-, as shown in the equation below.

$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$

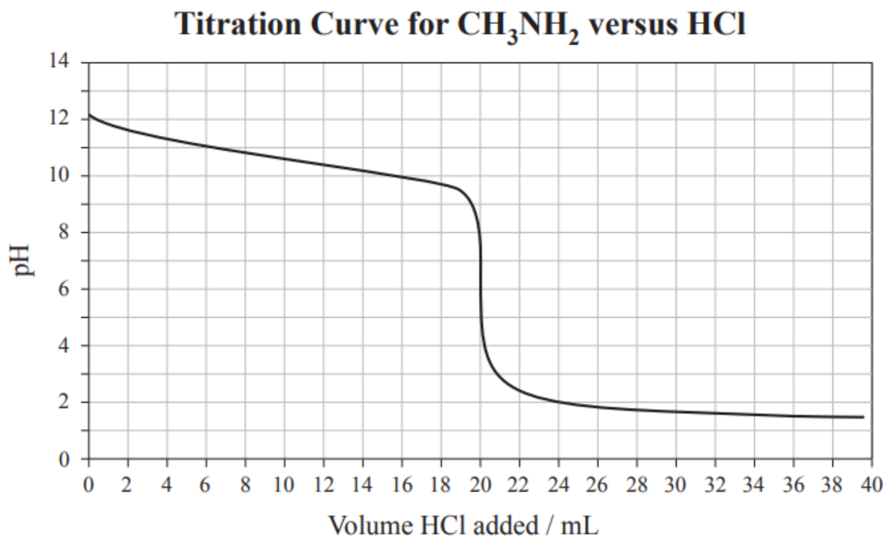
pK_a (CH_3COOH) = 4.76   Ka (CH_3COOH) = 1.74×10^{-5}

Calculate the pH of a 0.420 mol L^{-1} CH_3COONa solution.

(2018:2)

(b) (ii) A titration was carried out by adding 0.210 mol L^{-1} hydrochloric acid, HCl, to 25.0 mL of 0.168 mol L^{-1} methanamine, CH_3NH_2. The equation for the reaction is: $HCl + CH_3NH_2 \rightarrow CH_3NH_3^+ + Cl^-$

pK_a ($CH_3NH_3^+$) = 10.6 K_a($CH_3NH_3^+$) = 2.51×10^{-11}



Calculate the pH at the equivalence point.

(2017:2)

(a) Ammonia, NH_3, is a weak base. pK_a (NH_4^+) = 9.24 K_a (NH_4^+) = 5.75×10^{-10}

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

(2017:3)

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: $CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$

K_a (CH_3COOH) = 1.74×10^{-5}

(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}.

(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).

ANSWERS

(2018:1)

(a) $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$

$$H_3O^+ = \sqrt{\frac{K_a \times K_w}{[CH_3COO^-]}}$$
$$= \sqrt{\left(\frac{1.74 \times 10^{-5} \times 1 \times 10^{-14}}{0.420}\right)}$$
$$= 6.44 \times 10^{-10}$$

pH = $-\log 6.44 \times 10^{-10}$ = 9.19

(2018:2)

(b) (ii) $CH_3NH_3^+ + H_2O \rightleftharpoons CH_3NH_2 + H_3O^+$

$$K_a = 2.51 \times 10^{-11} = \frac{[H_3O^+]^2}{\left(0.168 \times \frac{25}{45}\right)}$$

$[H_3O^+] = 1.53 \times 10^{-6}$ mol L^{-1}
pH = $-\log 1.53 \times 10^{-6}$ = 5.82

(2017:2)

(a) (i) $[H_3O^+] = \sqrt{(Ka \times Kw \div [NH_3])}$
$[H_3O^+ = \sqrt{(5.75 \times 10^{-10} \times 1.00 \times 10^{-14} \div 0.105)}$
$[H_3O^+] = 7.40 \times 10^{-12}$ $molL^{-1}$ pH = $-\log (7.40 \times 10^{-12})$ = 11.1

(2017:3)

(b) (i) $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$ $K_a = [CH_3COO^-][H_3O^+] / [CH_3COOH]$
$1.74 \times 10^{-5} = (10^{-2.77})^2 / [CH_3COOH]$
$[CH_3COOH] = 0.166$ mol L^{-1}

(ii) n(NaOH) added = 0.112 mol L–1 × 0.01 L = 1.12×10^{-3} mol
Initial $n(CH_3COOH) = 0.166$ mol $L^{-1} \times 0.02$ L = 3.32×10^{-3} mol
$n(CH_3COOH)$ remaining in 30 mL = 3.32×10^{-3} mol – 1.12×10^{-3} mol = 2.2×10^{-3} mol

n(CH_3COO^- in 30 mL = 1.12×10^{-3} mol

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

$$1.74 \times 10^{-5} = \frac{\left(\dfrac{1.12 \times 10^{-3}}{0.03}\right)[H_3O^+]}{\dfrac{2.2 \times 10^{-3}}{0.03}}$$

$[H_3O^+] = 3.42 \times 10^{-5}$ pH = $-\log 3.42 \times 10^{-5}$ pH = 4.47