1. (a) Calculate the pH of a solution of ethanoic acid of concentration 0·01 mol l^{-1}.

(b) A mixture of ethanoic acid and sodium ethanoate solution constitutes a buffer solution. Show how this buffer solution is able to resist a change in pH when small quantities of the following are added:

(i) hydrochloric acid.

(ii) sodium hydroxide solution.

(Use of equations may be helpful).

(c) A buffer solution was prepared from equal volumes of 0·25 mol l^{-1} ethanoic acid and 0·15 mol l^{-1} sodium ethanoate. Calculate the pH of the buffer solution.

2. Aqueous ethanoic acid ionises as follows:

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

$\Delta H = 0$ kJ mol-1 Ka = $1{\cdot}8 \times 10^{-5}$

(a) Explain what will happen to Ka if the temperature is raised.

(b) A little hydrochloric acid is added to the ethanoic acid solution.

(i) What will be the effect on the Ka? Explain your reasoning.

(ii) What will be the effect on the concentration of ethanoate ions in the solution. Explain your answer.

3. Benzoic acid (C_6H_5COOH) is a weak monobasic acid (Ka = $6{\cdot}4 \times 10^{-5}$)

(a) Calculate the hydrogen in concentration in 0·02 mol l^{-1} benzoic acid.

(b) What is the pH of 0·02 mol l^{-1} benzoic acid?

(c) What is the pH of a solution containing 7·2g of sodium benzoate in 1 litre of 0·02 mol l^{-1} benzoic acid?

Ex 2.12

1 (a) $pH = \frac{1}{2} pK_a - \frac{1}{2} \log C$

$= \frac{1}{2} \times 4.77 - \frac{1}{2} \log(00)$

$= 2.385 - \frac{1}{2} \times (-2)$

$= \underline{3.38}$

$\left(K_{a\,(CH_3COOH)} = 1.7 \times 10^{-5} \quad pK_a = -\log(1.7 \times 10^{-5}) = \underline{4.77} \right)$

(b) (i) When H^+ ions are added the weak acid equilibrium is affected. The equilibrium will shift to reduce the effect of the increase in H^+ concentration. This will occur by the ethanoate ions joining with the H^+ ions added to form ethanoic acid molecules

$$H^+_{(aq)} + CH_3COO^-_{(aq)} \rightarrow CH_3COOH_{(l)}$$

(ii) When hydroxide ions are added they react with H^+ ion making water. This decrease the concentration of the H^+ ions which will affect the weak acid equilibrium. The equilibrium will shift to reduce the decrease in H^+ ions. It will do this by ethanoic acid molecules ionising to make more H^+ ions

$$CH_3COOH_{(l)} \rightarrow CH_3COO^-_{(aq)} + H^+_{(aq)}$$

(c) $pH = pK_a - \log \frac{[acid]}{[Salt]}$

$= 4.77 - \log \frac{0.25}{0.15} = 4.77 - \log 1.66$

$= 4.77 - 0.22$

$= \underline{\underline{4.55}}$

2. (a) No effect as $\Delta H = 0$

(b) (I) No effect, add HCl the increase concentration of H^+ ions but Ka is independ-ent of changes in concentration. The equilibrium will shift to reduce concentration of H^+ ions but when equilibrium is re-established and the new concentrations of reactants products are use to calculate Ka, the same value is obtained.

(II) The concentration of ethanoate ions will decrease. This is because when the Hydrochloric acid is added the concentration of H^+ ion increases. This affects the weak acid equilibrium which will shift to reduce the H^+ ion concentration. It will do this by ethanoate ions joining with H^+ ion to form ethanoic acid molecules.

3 (a) $K_a = \frac{[H^+]^2}{c}$

$[H^+] = \sqrt{K_a \times c} = \sqrt{6.4\times10^{-5} \times 0.02}$

$= 0.00113$ mol/l

(b) $pH = -\log[H^+] = -\log(0.00113)$

$= \underline{2.95}$

(c) $pH = pK_a - \log\frac{[acid]}{[Salt]}$

$= 4.19 - \log\frac{0.02}{0.05}$

$= 4.1 (-0.4)$

$= \underline{4}$

$pK_a_{(C_6H_5COOH)} = -\log K_a = -\log(6.4\times10^{-5})$

$= \underline{4.19}$

$[C_6H_5COO^-] = \frac{mol}{vol} = \frac{0.059}{1}$

$= 0.05$ mol/l

moles $C_6H_5COO = \frac{7.2}{144} = 0.05$