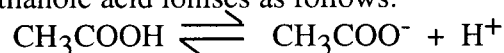


1. (a) Calculate the pH of a solution of ethanoic acid of concentration 0.01 mol l^{-1} .
 (b) A mixture of ethanoic acid and sodium ethanoate solution constitutes a buffer solution. Show how this buffer solution is able to resist a change in pH when small quantities of the following are added:
 - (i) 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:



$$\Delta H = 0 \text{ kJ mol}^{-1}$$

$$K_a = 1.8 \times 10^{-5}$$

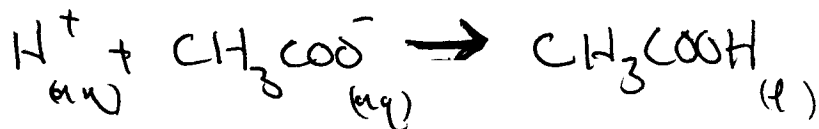
- (a) Explain what will happen to K_a if the temperature is raised.
 - (b) A little hydrochloric acid is added to the ethanoic acid solution.
 - (i) What will be the effect on the K_a ? Explain your reasoning.
 - (ii) What will be the effect on the concentration of ethanoate ions in the solution. Explain your answer.
3. Benzoic acid ($\text{C}_6\text{H}_5\text{COOH}$) is a weak monobasic acid ($K_a = 6.4 \times 10^{-5}$)
 - (a) Calculate the hydrogen ion 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)
$$\text{pH} = \frac{1}{2} \text{p}K_a - \frac{1}{2} \log C$$
$$= \frac{1}{2} \times 4.77 - \frac{1}{2} \log(0.01)$$
$$= 2.385 - \frac{1}{2} \times (-2)$$
$$= \underline{3.38}$$

$$K_a(\text{CH}_3\text{COOH}) = 1.7 \times 10^{-5}$$
$$\text{p}K_a = -\log(1.7 \times 10^{-5})$$
$$= \underline{4.77}$$

(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



(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

(c)
$$\text{CH}_3\text{COOH}_{(l)} \rightarrow \text{CH}_3\text{COO}^-_{(aq)} + \text{H}^+_{(aq)}$$

$$\text{pH} = \text{p}K_a - \log \frac{[\text{acid}]}{[\text{Salt}]}$$
$$= 4.77 - \log \frac{0.25}{0.15} = 4.77 - \log 1.66$$
$$= 4.77 - 0.22$$
$$= \underline{4.55}$$

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

(b) (i) No effect, add HCl to increase concentration of H^+ ions but K_a is independent 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 and products are used to calculate K_a , 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.

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

$$[H^+] = \sqrt{K_a \times C} = \sqrt{6.4 \times 10^{-5} \times 0.02}$$

$$= 0.00113 \text{ mol/l}$$

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

$$= \underline{2.95}$$

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

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

$$= 4.1 - (0.4)$$

$$= \underline{4}$$

$$pK_a = -\log K_a = -\log (6.4 \times 10^{-5})$$

$$= \underline{4.19}$$

$$[C_6H_8COO] = \frac{\text{mol}}{\text{vol}} = \frac{0.059}{1}$$

$$= 0.059 \text{ mol/l}$$

$$\text{moles } C_6H_8COO = \frac{7.2}{144} = 0.05$$