Problem 8-14 Buffers

Buffer consisting of 0.2 mol dm^{-3} of acetic acid and 0.2 mol dm^{-3} of sodium acetate has pH = 4.75. What change will cause the addition of 0.05 mol dm^{-3} KOH solution?

[$pH_1 = pK_{HAc} = 4.75$, $pH_2 = 4.97$]

Solution:

$HAc = H^+ + Ac^-$ conversion x $K_{HAc} = \frac{c_{H^+} \cdot c_{Ac^-}}{c_{HAc}}$

$NaAc = Na^+ + Ac^-$ completely dissociated

$c_{HAc} = 0.2$ mol dm^{-3}

$c_{NaAc} = c_{Ac-} = 0.2$ mol dm^{-3} (c_{Ac-} which is formed by dissociation of the acid is negligible due to very small value of the dissociation constant)

$$c_{H^+} = K_{HAc} \cdot \frac{c_{HAc}}{c_{Ac^-}} = K_{HAc} \cdot \frac{0.2}{0.2} = K_{HAc}$$

$pH_1 = pK_{HAc} = 4.75$

On addition of the strong base a reaction

$$HAc + KOH = KAc + H_2O$$

will occur and 0.05 mol dm^{-3} of acid HAc converts to 0.05 mol dm^{-3} of sodium acetate

$c_{HAc} = 0.2 - 0.05 = 0.15$ mol dm^{-3}

$c_{Ac-} = 0.2 + 0.05 = 0.25$

$$c_{H^+} = K_{HAc} \cdot \frac{c_{HAc}}{c_{Ac^-}} = 1.78 \cdot 10^{-5} \cdot \frac{0.15}{0.25} = 1.068 \cdot 10^{-5} \text{ mol dm}^{-3}$$

$pH_2 = -\log (1.068 \cdot 10^{-5}) = 4.9714$

pH increases by $\Delta pH = 4.9714 - 4.75 = 0.2214$