Answers to exam questions for Chapter 12 Buffers and titration curves

Question 1

Lactic acid

a pH = 3.86. A (pH = pK_a at this point)

b 20.0 mL A {n(HLac) = n(NaOH); c(HLac) = c(NaOH)}

c $\text{HLac} + \text{NaOH} \rightarrow \text{LacNa} + H_2O$ (reaction goes to completion)

$\text{Lac}^- + H_2O \rightleftharpoons \text{HLac} + OH^-$ (lactate ion reacts; basic solution)

$$K_b = \frac{[\text{HLac}][OH^-]}{[\text{Lac}^-]}$$

($K_b = K_w/K_a$; [HLac] = [OH^-])

$$7.25 \times 10^{-11} = \frac{[OH^-]^2}{0.0500 \text{ mol L}^{-1}}$$

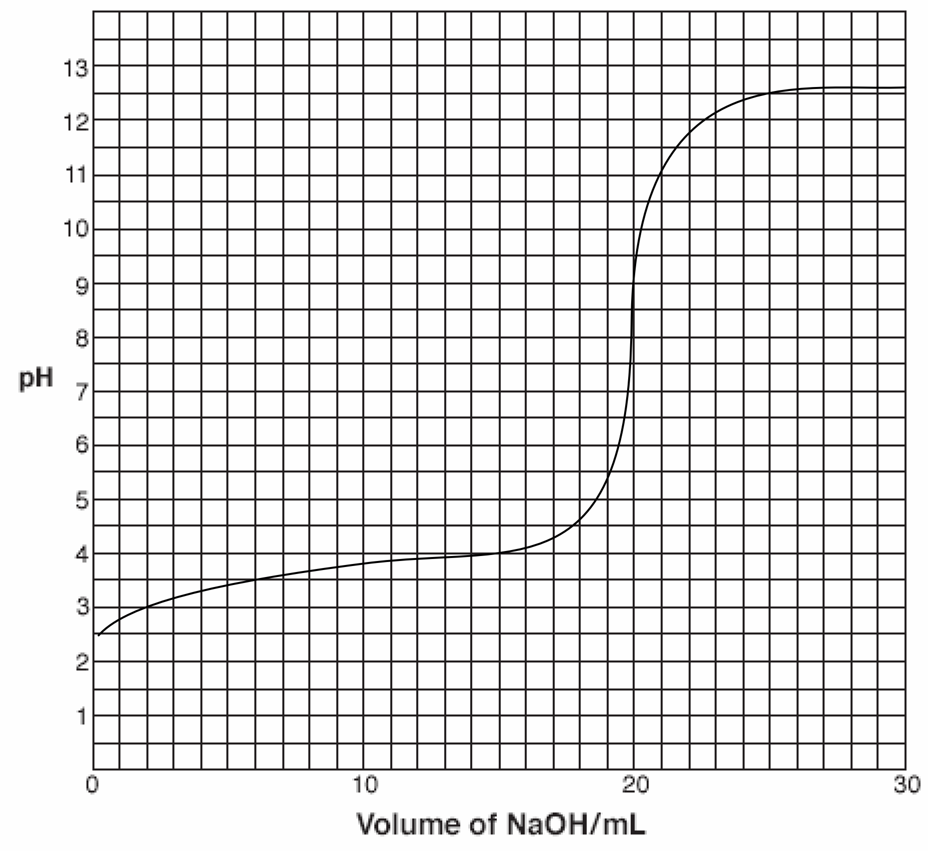
([Lac^-] = 0.0500 mol L^{-1}, volume doubled)

$$[OH^-] = \sqrt{7.25 \times 10^{-11} \times 0.0500 \text{ mol L}^{-1}}$$
$$= 1.90 \times 10^{-6} \text{ mol L}^{-1}$$

pOH = 5.72 (pOH = − $\log_{10}[OH^-]$)

pH = 8.28 (pH = 8.3) (pH = 14 − pOH)

M = correct method used with minor error, E = correct value



Put in start pH =2.43; at 10 mL pH =3.86; at 20 mL pH = 8.3; shape of WA *v* SB – curve at start and flat at end.

A = rough sketch with points not drawn accurately, M = better sketch and most points, E = correct shape and points correctly plotted

e D A (pK_a of indicator on vertical part of graph, colour change there)

Question 2

a A buffer solution is one that resists changes to its pH when moderate amounts of acid or a base are added. A (the definition of a buffer must be known)

b **i** $HCO_3^-(aq) + H_3O^+(aq) \rightarrow H_2CO_3(aq) + H_2O(l)$ A

ii $H_2CO_3(aq) + OH^-(aq) \rightarrow HCO_3^-(aq) + H_2O(l)$ A

c $H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$ (equilibrium equation)

$$K_a = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]}$$ (equilibrium expression)

$$\frac{[K_a]}{[H_3O^+]} = \frac{[HCO_3^-]}{[H_2CO_3]}$$ (rearrange expression)

$K_a(H_2CO_3) = 4.5 \times 10^{-7}$; pH = 6.11 so $[H_3O^+] = 7.76 \times 10^{-6}$ mol L^{-1} (substitute)

$$\frac{[HCO_3^-]}{[H_2CO_3]} = \frac{4.5 \times 10^{-7}}{7.76 \times 10^{-7}}$$

$$= 0.58$$

A = correct expression used, M = correct working and answer

Question 3

a The number that is closest to the pK_a value for the $H_3PO_4/H_2PO_4^-$ acid-base conjugate pair is 2.1. A
Justification: At half the equivalence point pH = pK_a because $[H_3PO_4] = [H_2PO_4^-]$ M

(from the equilibrium expression: $K_a = \frac{[H_2PO_4^-][H_3O^+]}{[H_3PO_4]}$; $K_a = [H_3O^+]$)

b **i** as an acid: $H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+$ A (proton donor)

ii as a base: $H_2PO_4^- + H_2O \rightleftharpoons H_3PO_4 + OH^-$ A (proton acceptor)

c The acid strength of $H_2PO_4^-$ is higher than its base strength. A
Justification: At the equivalence point the pH is below 7, the solution is acidic, hence the equilibrium in **b i** dominates over the equilibrium in **b ii**. M ($[H_2PO_4^-]$ is amphiprotic, and here its ability to act as an acid is greater than its ability to act as a base)

d bromocresol green (pK_a = 4.7) A (pH at equivalence point is 4.6; indicator is on vertical)

e The volume of NaOH that has been added at the equivalence point is 10.0 mL. A
{n(NaOH) = $n(H_3PO_4)$; c(NaOH) = $c(H_3PO_4)$ and ratio is 1:1}

f Volume at equivalence point is 20.0 mL

$$[H_2PO_4^-] = \frac{0.100 \text{ mol L}^{-1} \times 10 \text{ mL}}{20 \text{ mL}}$$

$$= 0.0500 \text{ mol L}^{-1}$$

M (concentration is halved as volume is doubled)