Marks 6

- The $H_2PO_4^-$ and HPO_4^{2-} ions play a major role in maintaining the intracellular pH balance. Write balanced equations to show how a solution containing these ions can act as a buffer.

The presence of $H_2PO_4^-(aq)$ means that the solution can remove $OH^-(aq)$:

$$H_2PO_4^-(aq) + OH^-(aq) \rightleftharpoons HPO_4^{2-}(aq) + H_2O(l)$$

The presence of $HPO_4^{2-}(aq)$ means that the solution can remove $H^+(aq)$:

$$HPO_4^{2-}(aq) + H^+(aq) \rightleftharpoons H_2PO_4^-(aq)$$

For phosphoric acid, $K_{a1} = 7.1 \times 10^{-3}$ M, $K_{a2} = 6.3 \times 10^{-8}$ M, $K_{a3} = 4.2 \times 10^{-13}$ M. At what pH would the $H_2PO_4^-$ / HPO_4^{2-} buffer system be most effective? Why?

Buffers are most effective when [acid] = [base] at which point pH = pK_a. For this system, this requires $[H_2PO_4^-(aq)] = [HPO_4^{2-}(aq)]$. This acid / base equilibrium corresponds to K_{a2}:

$$H_2PO_4^-(aq) \rightleftharpoons HPO_4^{2-}(aq) + H^+(aq)$$

Thus, pH = $pK_{a2} = -\log_{10}(6.3 \times 10^{-8}) = 7.20$

Calculate the ratio of $H_2PO_4^-$ / HPO_4^{2-} needed to give a solution buffered to a pH of 7.35.

Using the Henderson-Hasselbalch equation:

$$\text{pH} = pK_a + \log_{10}\left(\frac{[\text{base}]}{[\text{acid}]}\right) \text{ or } \log_{10}\left(\frac{[\text{base}]}{[\text{acid}]}\right) = \text{pH} - pK_a$$

Thus:

$$\log_{10}\left(\frac{[\text{base}]}{[\text{acid}]}\right) = \text{pH} - pK_{a2} = 7.35 - 7.20 = 0.15$$

$$\left(\frac{[\text{base}]}{[\text{acid}]}\right) = 10^{0.15} = 1.4 \text{ or } \left(\frac{[\text{acid}]}{[\text{base}]}\right) = \frac{1}{1.4} = 0.71$$