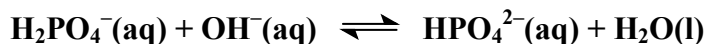


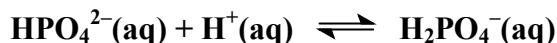
**Marks**  
**6**

- The  $\text{H}_2\text{PO}_4^-$  and  $\text{HPO}_4^{2-}$  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  $\text{H}_2\text{PO}_4^-(\text{aq})$  means that the solution can remove  $\text{OH}^-(\text{aq})$ :**



**The presence of  $\text{HPO}_4^{2-}(\text{aq})$  means that the solution can remove  $\text{H}^+(\text{aq})$ :**



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  $\text{H}_2\text{PO}_4^- / \text{HPO}_4^{2-}$  buffer system be most effective? Why?

**Buffers are most effective when  $[\text{acid}] = [\text{base}]$  at which point  $\text{pH} = \text{p}K_a$ . For this system, this requires  $[\text{H}_2\text{PO}_4^-(\text{aq})] = [\text{HPO}_4^{2-}(\text{aq})]$ . This acid / base equilibrium corresponds to  $K_{a2}$ :**



**Thus,  $\text{pH} = \text{p}K_{a2} = -\log_{10}(6.3 \times 10^{-8}) = 7.20$**

Calculate the ratio of  $\text{H}_2\text{PO}_4^- / \text{HPO}_4^{2-}$  needed to give a solution buffered to a pH of 7.35.

**Using the Henderson-Hasselbalch equation:**

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

**Thus:**

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

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