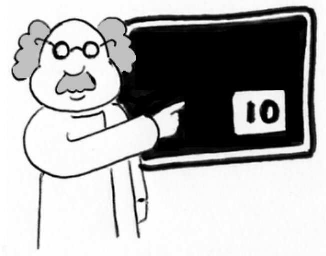

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3.5. Buffer solutions

A buffer solution is a solution that resists a change in pH when a small quantity of acid or base is added.

1. (a) A buffer solution is made by mixing 0.510 mol of methanoic acid with 0.450 mol of sodium methanoate in 500 cm^3 of water.

(i) Write an equation to represent the equilibrium established in the buffer solution.

..........(1 mark)

(ii) Calculate the pH of the buffer solution formed. (pK_a for methanoic acid = 3.75)

....

....

..........(3 marks)

(b) Explain how this buffer resists change in pH on;

(i) addition of a small quantity of acid.

....

..........(1 mark)

(ii) addition of a small quantity of base.

....

..........(1 mark)

2. Mark and Karen are carrying out a science project on the application of buffer solutions in the human body. They have discovered that a buffer of carbonic acid (H_2CO_3) and hydrogen carbonate (HCO_3^-) is present in blood plasma to maintain a pH of between 7.35 and 7.45.

(a) They would like to recreate a similar buffer solution in the laboratory. In what proportions should they mix 0.150 mol dm^{-3} solutions of carbonic acid and sodium hydrogen carbonate to give a buffer solution with a pH of 7.40? (K_a for H_2CO_3 is 4.5×10^{-7} mol dm^{-3}).

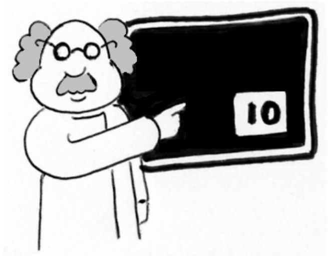
....

....

..........(2 marks)

(b) Why do you think buffer solutions are needed in the human body?

..........(2 marks)



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3. Acids and bases answers

3.5. Buffer solutions

1. (a) (i) $HCOOH(aq) \rightleftharpoons HCOO^-(aq) + H^+(aq)$ *(1 mark)*

(ii) $pK_a = -\log K_a$, $\therefore \underline{K_a = 10^{-3.75} = 1.78 \times 10^{-4}\ \text{mol dm}^{-3}}$ *(1 mark)*

$$K_a = \frac{[HCOO^-(aq)][H^+(aq)]}{[HCOOH(aq)]}$$

$[HCOO^-(aq)] = 0.450\ \text{mol} / 0.5\ \text{dm}^3 = 0.90\ \text{mol dm}^{-3}$

$[HCOOH(aq)] = 0.510\ \text{mol} / 0.5\ \text{dm}^3 = 1.02\ \text{mol dm}^{-3}$

Substituting these values in we get, $1.78 \times 10^{-4}\ \text{mol dm}^{-3} = 0.90 \times [H^+(aq)] / 1.02$

$\therefore [H^+(aq)] = \underline{2.02 \times 10^{-4}\ \text{mol dm}^{-3}}$ *(1 mark)*

$\therefore \underline{pH = 3.70}$ *(1 mark)*

(b) (i) On the addition of H^+ ions, according to Le Châtelier's principle, the equilibrium shifts to the left to remove the extra H^+ ions added and maintain the pH approximately constant. *(1 mark)*

(ii) On the addition of OH^- ions, the OH^- ions react with the HCOOH to produce water molecules and more $HCOO^-$;

$$HCOOH + OH^- \rightarrow HCOO^- + H_2O$$

This removes the OH^- and so the pH remains approximately constant. *(1 mark)*

2. (a) $H_2CO_3(aq) \rightleftharpoons HCO_3^-(aq) + H^+(aq)$

pH of desired buffer = 7.40, so $[H^+(aq)] = 10^{-7.40} = 3.98 \times 10^{-8}\ \text{mol dm}^{-3}$ *(1 mark)*

$$K_a = \frac{[HCO_3^-(aq)][H^+(aq)]}{[H_2CO_3(aq)]}$$

$$\therefore \frac{[HCO_3^-(aq)]}{[H_2CO_3(aq)]} = \frac{K_a}{[H^+(aq)]} = \frac{4.5 \times 10^{-7}\ \text{mol dm}^{-3}}{3.98 \times 10^{-8}\ \text{mol dm}^{-3}} = \frac{11.3}{1}$$ *(1 mark)*

Since both stock solutions are of an equal concentration they should mix the two in a ratio of

$\underline{11.3 : 1\ HCO_3^- : H_2CO_3}$

(b) Many reactions in the human body rely on enzymes. Enzymes work only under very precise conditions. If the pH moves outside of a narrow range, the enzymes slow or stop working and can be denatured. Hence maintaining a constant pH is essential. *(2 marks)*

3.6. More complex buffer calculations

1. $CH_3CH_2COOH + NaOH \rightarrow CH_3CH_2COO^-Na^+ + H_2O$

Moles of NaOH = $0.015\ \text{dm}^3 \times 0.100\ \text{mol dm}^{-3} = 1.5 \times 10^{-3}$ mol *(1 mark)*

$\therefore$ moles of CH_3CH_2COOH will decrease by 1.5×10^{-3} mol and moles of $CH_3CH_2COO^-Na^+$ will increase by 1.5×10^{-3} mol. *(1 mark)*