

# STARTER FOR 10!!!

## 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<sup>3</sup> 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)

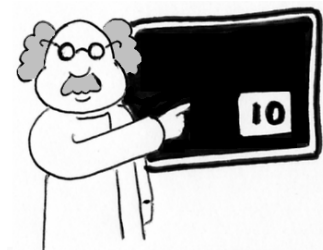
2. Mark and Karen are carrying out a science project on the application of buffer solutions in the human body. They have discovered that a buffer of carbonic acid (H<sub>2</sub>CO<sub>3</sub>) and hydrogen carbonate (HCO<sub>3</sub><sup>-</sup>) is present in blood plasma to maintain a pH of between 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<sup>-3</sup> solutions of carbonic acid and sodium hydrogen carbonate to give a buffer solution with a pH of 7.40? ( $K_a$  for H<sub>2</sub>CO<sub>3</sub> is  $4.5 \times 10^{-7}$  mol dm<sup>-3</sup>).

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

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

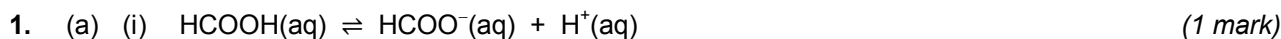
.....(2 marks)



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## 3. Acids and bases answers

### 3.5. Buffer solutions



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

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

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

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

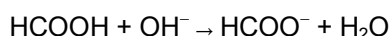
Substituting these values in we get,  $1.78 \times 10^{-4} \text{ mol dm}^{-3} = 0.90 \times [\text{H}^+(\text{aq})] / 1.02$

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

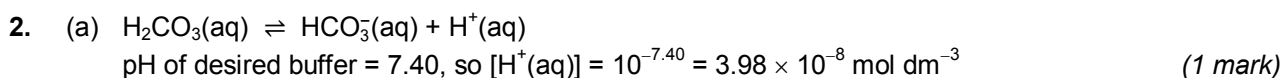
$$\therefore \text{pH} = \underline{3.70} \quad (1 \text{ mark})$$

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

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



This removes the  $\text{OH}^-$  and so the pH remains approximately constant. (1 mark)



$$K_a = \frac{[\text{HCO}_3^-(\text{aq})][\text{H}^+(\text{aq})]}{[\text{H}_2\text{CO}_3(\text{aq})]}$$

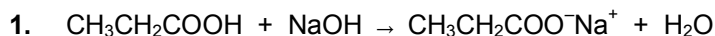
$$\therefore \frac{[\text{HCO}_3^-(\text{aq})]}{[\text{H}_2\text{CO}_3(\text{aq})]} = \frac{K_a}{[\text{H}^+(\text{aq})]} = \frac{4.5 \times 10^{-7} \text{ mol dm}^{-3}}{3.98 \times 10^{-8} \text{ mol dm}^{-3}} = \frac{11.3}{1} \quad (1 \text{ mark})$$

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

$$\underline{11.3 : 1 \text{ HCO}_3^- : \text{H}_2\text{CO}_3}$$

(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



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

$\therefore$  moles of  $\text{CH}_3\text{CH}_2\text{COOH}$  will decrease by  $1.5 \times 10^{-3} \text{ mol}$  and moles of  $\text{CH}_3\text{CH}_2\text{COO}^-\text{Na}^+$  will increase by  $1.5 \times 10^{-3} \text{ mol}$ . (1 mark)