Marks 7

- The pK_a of formic acid, HCO_2H, is 3.77. What is the pH of a 0.20 M solution of formic acid?

As formic acid is a weak acid, $[H_3O^+]$ must be calculated using a reaction table:

	HCO_2H	H_2O	$\rightleftharpoons$	H_3O^+	HCO_2^-
initial	**0.20**	**large**		**0**	**0**
change	$-x$	**negligible**		$+x$	$+x$
final	$0.20 - x$	**large**		x	x

The equilibrium constant K_a is given by:

$$K_a = \frac{[H_3O^+][HCO_2^-]}{[HCO_2H]} = \frac{x^2}{0.20 - x}$$

As pK_a = $-\log_{10}K_a$, $K_a = 10^{-3.77}$ and is very small, $0.20 - x \sim 0.20$ and hence:

$$x^2 = 0.20 \times 10^{-3.77} \quad \text{or} \quad x = 5.8 \times 10^{-3}\text{ M} = [H_3O^+]$$

Hence, the pH is given by:

$$\text{pH} = -\log_{10}[H_3O^+] = -\log_{10}(5.8 \times 10^{-3}) = 2.23$$

pH = **2.23**

Give the equation for the reaction of formic acid with solid sodium hydroxide.

$$HCOOH(aq) + NaOH(s) \rightarrow HCO_2^-(aq) + Na^+(aq) + H_2O(l)$$

Calculate the ratio of formate ion / formic acid required to give a buffer of pH 4.00.

Using the Henderson-Hasselbalch equation,

$$\text{pH} = \text{p}K_a + \log\frac{[\text{base}]}{[\text{acid}]}$$

$$4.00 = 3.77 + \log\frac{[HCO_2^-]}{[HCO_2H]}$$

$$\text{So, } \frac{[HCO_2^-]}{[HCO_2H]} = 10^{0.23} = 1.70$$

Answer: **1.70**

ANSWER CONTINUES ON THE NEXT PAGE

What amount (in mol) of sodium hydroxide must be added to 100.0 mL of 0.20 M HCO_2H to prepare a solution buffered at pH 4.00?

If the concentration of OH^- which is added is x M then this will react with HCO_2H to produce HCO_2^- so that:

$[HCO_2H] = (0.20 - x)$ M and
$[HCO_2^-] = x$ M

From above, if pH = 4.00, then $\frac{[HCO_2^-]}{[HCO_2H]} = 1.70$. Hence:

$$\frac{x}{0.20 - x} = 1.70 \quad \text{so } x = 0.13$$

To achieve $[OH^-(aq)] = 0.13$ mol L^{-1} in 100.0 mL, the number of moles of NaOH that must be added is:

number of moles = concentration × volume
= 0.13 mol L^{-1} × 0.1000 L = 0.013 mol

Answer: **0.013 mol**