Marks 4

- What is the pH of a 0.100 M solution of sodium acetate?
  The pK_a of acetic acid is 4.76.

Acetate is a weak base so [OH^-] must be calculated by considering the equilibrium:

	$CH_3CO_2^-$	H_2O	$\rightleftharpoons$	CH_3COOH	OH^-
initial	**0.100**	**large**		**0**	**0**
change	**-*x***	**negligible**		**+*x***	**+*x***
final	**0.100 – *x***	**large**		***x***	***x***

The equilibrium constant K_b is given by:

$$K_b = \frac{[CH_3COOH][OH^-]}{[CH_3CO_2^-]} = \frac{x^2}{(0.100-x)}$$

For an acid and its conjugate base:

pK_a + pK_b = 14.00

pK_b = 14.00 – 4.76 = 9.24

As pK_b = 9.24, $K_b = 10^{-9.24}$. K_b is very small so 0.100 – *x* ~ 0.100 and hence:

$x^2 = 0.100 \times 10^{-9.24}$ or $x = 7.59 \times 10^{-6}$ M = [OH^-]

Hence, the pOH is given by:

pOH = $-\log_{10}[OH^-] = \log_{10}[7.59 \times 10^{-6}] = 5.12$

Finally, pH + pOH = 14.00 so

pH = 14.00 – 5.12 = 8.88

pH = **8.88**

ANSWER CONTINUES ON THE NEXT PAGE

What is the ratio of acetate ion to acetic acid in this solution?

From the Henderson–Hasselbalch equation,

$$\mathbf{pH = pK_a + \log\frac{[base]}{[acid]} = 4.76 + \log\frac{[CH_3CO_2^-]}{[CH_3COOH]}}$$

At pH = 8.88,

$$\mathbf{8.88 = 4.76 + \log\frac{[CH_3CO_2^-]}{[CH_3COOH]}}$$

so

$$\mathbf{\frac{[CH_3CO_2^-]}{[CH_3COOH]} = 10^{4.12} = 1.3 \times 10^4}$$

Answer: $\mathbf{1.3 \times 10^4}$

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.