Marks

- What is the pH of a 0.020 M solution of HF? The pK_a of HF is 3.17. **2**

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

	HF	H_2O	$\rightleftharpoons$	H_3O^+	F^-
initial	**0.020**	**large**		**0**	**0**
change	**-x**	**negligible**		**+x**	**+x**
final	**0.020 – x**	**large**		**x**	**x**

The equilibrium constant K_a is given by:

$$K_a = \frac{[H_3O^+][F^-]}{[HF]} = \frac{x^2}{0.020 - x}$$

As pK_a = -log$_{10}$(K_a), $K_a = 10^{-3.17}$ and is very small. Hence, 0.020 – x ~ 0.020 and hence:

$$x^2 = 0.020 \times 10^{-3.17} \quad \text{or} \quad x = 3.68 \times 10^{-3}\ M = [H_3O^+]$$

Hence, the pH is given by:

$$pH = -\log_{10}[H_3O^+] = -\log_{10}[(3.68 \times 10^{-3})] = 2.43$$

pH = **2.43**

- What is the pH of a solution that is 0.075 M in acetic acid and 0.150 M in sodium acetate? The pK_a of CH_3COOH is 4.76. **2**

This solution containing a weak acid (CH_3COOH) and its conjugate base (CH_3COO^-) is a buffer and the Henderson-Hasselbalch equation can be used to calculate the pH:

$$pH = pK_a + \log_{10}\left(\frac{[base]}{[acid]}\right) = pK_a(CH_3COOH) + \log_{10}\left(\frac{[CH_3COO^-]}{[CH_3COOH]}\right)$$

$$pH = 4.76 + \log_{10}\left(\frac{0.150}{0.075}\right) = 5.06$$

pH = **5.06**

THE ANSWER CONTINUES ON THE NEXT PAGE

- What is the pH of a 0.010 M solution of $Ba(OH)_2$?

2

$Ba(OH)_2$ is a strong base and dissociates completely according to the equation:

$Ba(OH)_2(aq) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$

Hence a 0.010 M solution has $[OH^-(aq)] = 2 \times 0.010 = 0.020$ M.

As $pOH = -\log_{10}([OH^-(aq)]$, $pOH = -\log_{10}(0.020) = 1.70$.

In aqueous solution, $pH + pOH = 14.00$ so $pH = (14.00 - 1.70) = 12.30$

pH = **12.30**