Marks

- Calculate the pH of a 0.020 M solution of $Ba(OH)_2$. **1**

$Ba(OH)_2$ is a strong base so it will completely dissociate in solution:

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

As each $Ba(OH)_2$ dissociates to make $2OH^-$, a 0.020 M solution has $[OH^-(aq)]$ = 0.040 M.

By definition, pOH = $-\log_{10}[OH^-(aq)]$ so pOH = $-\log_{10}(0.040)$ = 1.40.

As pH + pOH = 14.00,

pH = 14.00 – 1.40 = 12.60

pH = **12.60**

- Calculate the pH of a 0.150 M solution of HNO_2. The pK_a of HNO_2 is 3.15. **3**

As HNO_2 is a weak acid, $[H_3O^+]$ must be calculated using a reaction table:

	HNO_2	H_2O	$\rightleftharpoons$	H_3O^+	NO_2^-
initial	**0.150**	**large**		**0**	**0**
change	$-x$	**negligible**		$+x$	$+x$
final	$0.150 - x$	**large**		x	x

The equilibrium constant K_a is given by:

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = \frac{x^2}{0.150 - x}$$

As pK_a = $-\log_{10}K_a$, $K_a = 10^{-3.15}$ and is very small, $0.150 - x \sim 0.150$ and hence:

$$x^2 = 0.150 \times 10^{-3.15} \quad \text{or} \quad x = 1.03 \times 10^{-2}\ \text{M} = [H_3O^+]$$

Hence, the pH is given by:

$$\text{pH} = -\log_{10}[H_3O^+] = -\log_{10}(1.03 \times 10^{-2}) = 1.987$$

pH = **1.987**

ANSWER CONTINUES ON THE NEXT PAGE

- Calculate the pH of a solution that is 0.080 M in acetic acid and 0.160 M in sodium acetate. The pKa of acetic acid is 4.76.

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The solution contains both a weak acid (acetic acid) and its conjugate base (the acetate ion). This is a buffer and its pH can be calculate using the Henderson-Hasselbalch equation. With [acid] = 0.080 M and [base] = 0.160 M,

$$\mathbf{pH = p}\boldsymbol{K}_\mathbf{a} + \mathbf{log}\left(\frac{[\mathbf{base}]}{[\mathbf{acid}]}\right) = \mathbf{4.76 + log}\left(\frac{\mathbf{0.160}}{\mathbf{0.080}}\right) = \mathbf{5.06}$$

pH = **5.06**