Marks 8

- Solution A consists of a 0.15 M aqueous solution of nitrous acid (HNO_2) at 25 °C. Calculate the pH of Solution A. The pK_a of HNO_2 is 3.15.

Nitrous acid is a weak acid so $[H_3O^+]$ must again be calculated:

	HNO_2(aq)	H_2O(l)	$\rightleftharpoons$	H_3O^+(aq)	NO_2^-(aq)
initial	0.15	large		0	0
change	$-x$	negligible		$+x$	$+x$
final	$0.15 - x$	large		x	x

The equilibrium constant K_a is given by:

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = \frac{x^2}{0.15 - x}$$

As $K_a = 10^{-3.15}$ is very small, $0.15 - x \sim 0.15$ and hence:

$$x^2 = 0.15 \times 10^{-3.15} \quad \text{or} \quad x = 0.0103 \text{ M} = [H_3O^+(aq)]$$

Hence, the pH is given by:

$$pH = -\log_{10}[H_3O^+(aq)] = -\log_{10}[0.0103] = 1.99$$

ANSWER: **pH = 1.99**

ANSWER CONTINUES ON THE NEXT PAGE

At 25 °C, 1.00 L of Solution B consists of 13.8 g of sodium nitrite ($NaNO_2$) dissolved in water. Calculate the pH of Solution B.

The formula mass of $NaNO_2$ is (22.99 (Na) + 14.01 (N) + 2×16.00 (O)) g mol^{-1} = 69 g mol^{1}. 13.8 g therefore corresponds to:

$$\text{amount of } NaNO_2 = \frac{\text{mass}}{\text{formula mass}} = \frac{13.8\,\text{g}}{69.0\,\text{g mol}^{-1}} = 0.200\,\text{mol}$$

A 1.00 L solution containing this amount has a molarity of 0.200 M. The nitrite ion acts as a base and [OH^-(aq)] must be calculated from the equilibrium:

	NO_2^-(aq)	H_2O(l)	$\rightleftharpoons$	OH^-(aq)	HNO_2(aq)
initial	0.200	large		0	0
change	$-y$	negligible		$+y$	$+y$
final	$0.200 - y$	large		y	y

The equilibrium constant K_b is given by:

$$K_a = \frac{[OH^-][HNO_2]}{[NO_2^-]} = \frac{y^2}{0.200 - y}$$

In aqueous solution, pK_a + pK_b = 14.00. Hence pK_b = (14.00 – 3.15) = 10.85 and as $K_b = 10^{-10.85}$ is very small, $0.200 - y \sim 0.200$ and hence:

$$y^2 = 0.200 \times 10^{-10.85} \quad \text{or} \quad y = 1.68 \times 10^{-6}\text{ M} = [OH^-(aq)]$$

Hence, the pOH is given by:

$$\text{pOH} = -\log_{10}[OH^-(aq)] = -\log_{10}[1.68 \times 10^{-6}] = 5.77$$

As pH + pOH = 14, pH = 8.23

ANSWER: **pH = 8.23**

ANSWER CONTINUES ON THE NEXT PAGE

Solution B (1.00 L) is poured into Solution A (1.00 L) and allowed to equilibrate at 25 °C. Calculate the pH of the final solution.

This solution contains an acid and its conjugate base so the Henderson-Hasselbalch equation can be used. As [acid] = [HNO_2] = 0.15 M and [base] = [NO_2^-] = 0.200 M:

$$\mathbf{pH = p}\boldsymbol{K}_\mathbf{a} + \mathbf{log_{10}}\left(\frac{\mathbf{[base]}}{\mathbf{[acid]}}\right) = \mathbf{3.15 + log_{10}}\left(\frac{\mathbf{0.200}}{\mathbf{0.15}}\right) = \mathbf{3.27}$$

ANSWER: **pH = 3.27**

If you wanted to adjust the pH of the mixture of Solution A and Solution B to be exactly equal to 3.00, which component in the mixture would you need to increase in concentration?

The acid, HNO_2