

Solutions

Original marking guidance: M1 method, A1 accuracy, B1 explanation.

2.1 NH_4^+ ; HPO_4^{2-} .

2.2 $\text{pH} = -\log_{10}(0.0250) = 1.60$.

2.3 $[\text{H}^+] = K_w/[\text{OH}^-] = 5.00 \times 10^{-12} \text{ mol dm}^{-3}$; $\text{pH} = 11.30$.

3.1 $x = 10^{-3.00} = 0.00100 \text{ mol dm}^{-3}$. $K_a = x^2/(0.0800 - x) = 1.27 \times 10^{-5} \text{ mol dm}^{-3}$. Percentage = $100(0.00100/0.0800) = 1.25\%$.

3.2 Weak acid $[\text{H}^+]$ falls by $\sqrt{10}$, giving pH increase about 0.5. Strong acid pH increases by 1. At extreme dilution water ionisation is significant, or increased dissociation can invalidate the weak-acid approximation.

3.3 $n = m/M = 0.610/122 = 0.00500 \text{ mol}$; $c = n/V = 0.00500/0.2500 = 0.0200 \text{ mol dm}^{-3}$. $x = 10^{-3.00} = 0.00100 \text{ mol dm}^{-3}$. $K_a = x^2/(c - x) = 10^{-6}/0.0190 = 5.26 \times 10^{-5} \text{ mol dm}^{-3}$. Calibrate the pH meter with suitable buffers and control temperature, or use a volumetric flask accurately/rinse the electrode between samples.

3.4 Partial dissociation gives the weak acid fewer H^+ ions and higher pH than the equimolar strong acid. The weak base gives fewer OH^- ions and lower pH than the strong alkali. Ethanoate accepts H^+ from water, producing OH^- , so its salt is alkaline. Strong-acid pH rises by 2 and 3 respectively. Small-dissociation weak-acid pH rises approximately by 1 and 1.5. These comparisons require the approximations to remain valid; water contribution or increased fractional dissociation can invalidate extrapolation.