

Report Table (B) B: Part C Data: pH Measurements of Dilute Solutions of Acetic Acid

Solution	pH using pH paper
0.10 M $\text{HC}_2\text{H}_3\text{O}_2$	9.0 0.1 0.2 0.3 0.4 0.5 0.6 0.7 0.8 0.9 1.0
0.010 M $\text{HC}_2\text{H}_3\text{O}_2$	3.4 3.5 3.6 3.7 3.8 3.9 4.0 4.1 4.2 4.3 4.4
0.0010 M $\text{HC}_2\text{H}_3\text{O}_2$	5.0 5.1 5.2 5.3 5.4 5.5 5.6 5.7 5.8 5.9 6.0

Write the net ionic equation describing the dissociation of acetic acid in aqueous solution.

Calculate $[\text{H}^+]$ and $[\text{C}_2\text{H}_3\text{O}_2^-]$ for each concentration above, showing a sample calculation here.

$$\begin{aligned}
 \text{pH} &= -\log [\text{H}^+] & [\text{H}_2\text{PO}_4^-] &= 0.10 \text{ M} \\
 9.0 &= -\log x & x &= [0.10 - 0.0001] \text{ M} = 0.0999 \text{ M} \\
 -\log x &= 9.0 & K &= \frac{[0.0001 \text{ M}][0.0001 \text{ M}]}{[0.0999 \text{ M}]} \\
 x &= 10^{-9.0} & &= \frac{0.0001 \text{ M}}{0.0999 \text{ M}} \\
 x &= 0.0001 \text{ M} & &= 0.001 \text{ M}
 \end{aligned}$$

Write the equilibrium constant expression for acetic acid and calculate a K value and an apparent percent dissociation for each concentration, showing sample calculations here.

$$\begin{aligned}
 K_a &= \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = \frac{x^2}{0.10 - x} = \frac{(1.0 \times 10^{-9})^2}{0.10 - 1.0 \times 10^{-9}} = 1.0 \times 10^{-17} \\
 \text{Percent dissociation} &= \frac{[\text{H}^+]}{[\text{HC}_2\text{H}_3\text{O}_2]_{\text{initial}}} \times 100 = \frac{1.0 \times 10^{-9}}{0.10} \times 100 = 1.0 \times 10^{-7} \%
 \end{aligned}$$

What is the effect of dilution on the percent dissociation? Why should this behavior be expected?

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