At 25 °C, 0.10 mol each of a strong acid and OH\(^-\) are combined in 1.0 L of water. The pH of the solution is...

25% 1. < 7
25% 2. = 7
25% 3. > 7
25% 4. Further information needed

Step 2: [H\(_3\)O\(^+\)] when "too much" base added

\[ \text{[H}_3\text{O}^+] \text{ when "too much" base added} \]

\[ K_w = 200, \text{ mL of } c_b = 0.30 \text{ M of OH}^- \text{ is combined with } V_a = 200, \text{ mL of } c_a = 0.20 \text{ M of HA, } K_a = 1.0 \times 10^{-6} \text{ and } K_b = K_w / K_a = 1.0 \times 10^{-8}. \]
Step 2: $[\text{H}_3\text{O}^+]$ when “too much” base added

\[ K_c = 200. \text{ mL of } c_b = 0.30 \text{ M of OH}^- \text{ is combined with } V_a = 200. \text{ mL of} \]
\[ c_a = 0.20 \text{ M of HA, } K_a = 1.0 \times 10^{-6} \text{ and } K_b = \frac{K_w}{K_a} = 1.0 \times 10^{-8}. \]

Step 1 results:

Ignoring acid-base equilibrium, after the limiting reagent reaction is complete, the resulting concentrations of HA, $A^-$, and OH$^-$ are ...

\[ [\text{HA}] = 0 \]
\[ [A^-] = \frac{(c_a V_a)}{(V_a + V_b)} = \frac{(0.040 \text{ mol})}{(0.400 \text{ L})} = 0.10 \text{ M} \]
\[ [\text{OH}^-] = \frac{(c_b V_b - c_a V_a)}{(V_a + V_b)} = \frac{(0.060 - 0.040) \text{ mol}}{(0.400 \text{ L})} = 0.020 \text{ mol} / (0.400 \text{ L}) = 0.050 \text{ M} \]

Step 2: $[\text{H}_3\text{O}^+]$ when “too much” base added

If conjugate base is present and excess $[\text{OH}^-]$, then work with

\[ \text{H}_2\text{O}(l) + A^- (aq) \rightleftharpoons \text{HA}(aq) + \text{OH}^- (aq), K_a = K_w / K_a \]

\[ \text{Initial} \]

\begin{array}{cccc}
\text{A}^- (aq) & \text{HA}(aq) & \text{OH}^- (aq) & Q \\
0 & 0.10 & 0.050 & 0 \\
\end{array}

\[ \text{Change} \]

\begin{array}{cccc}
x & x & x & 1.0 \times 10^{-8} \\
\end{array}

\[ \text{Equilibrium} \]

\begin{array}{cccc}
0.10 - x & x & 0.050 + x & 1.0 \times 10^{-8} \\
\end{array}

\[ \text{Approximate} \]

\begin{array}{cccc}
0.10 & x & 0.050 & 1.0 \times 10^{-8} \\
\end{array}

\[ [\text{OH}^-] = 0.050 \text{ M (easy!)} \]
\[ [\text{H}_3\text{O}^+] = K_a / [\text{OH}^-] = (1.0 \times 10^{-14}) / (0.050) = 2.0 \times 10^{-13} \]
\[ [\text{HA}] = x = K_a / [A^-] / [\text{OH}^-] = 1.0 \times 10^{-8} \times 0.10 / 0.050 = 2.0 \times 10^{-8} \text{ (tiny!)} \]
Too little, just enough, too much?

[TP] At 25 °C, 0.10 mol each of a strong acid and OH⁻ are combined in 1.0 L of water. The pH of the solution is ...

25% 1. < 7
25% 2. = 7
25% 3. > 7
25% 4. Further information needed

[Quiz] At 25 °C, 0.10 mol each of a weak acid HA and its conjugate base A⁻, and OH⁻ are combined in 1.0 L of water. The pH of the solution is ...

25% 1. < 7
25% 2. = 7
25% 3. > 7
25% 4. Further information needed

[Quiz] At 25 °C, 0.10 mol each of a weak acid HA and a strong acid HB, and 0.20 mol of OH⁻ are combined in 1.0 L of water. The pH of the solution is ...

25% 1. < 7
25% 2. = 7
25% 3. > 7
25% 4. Further information needed
At 25 °C, 0.10 mol each of a weak acid HA, a strong acid HB, and OH⁻ are combined in 1.0 L of water. The pH of the solution is ...

25% 1. < 7
25% 2. = 7
25% 3. > 7
25% 4. Further information needed

An acid HA has $K_a = 1.0 \times 10^{-4}$ at 25 °C. An aqueous solution is made by combining of 0.137 moles each of HA and A⁻ in a total volume of 250 mL at 25 °C. The pH of the solution is ...

17% 1. 1
17% 2. 4
17% 3. 7
17% 4. 10
17% 5. 13
17% 6. More information needed

At 25 °C, the pH of a 1.0 L solution of $c_a = c_b = 1.00 M, K_a = 1 \times 10^{-5}$ is ...

pH = 5.00

Add 100. mL of 0.100 M NaOH ...

$4HA(aq) + OH^- (aq) \rightarrow H_2O(l) + 1A^- (aq)$

HA → 1.00 mol - 0.010 mol = 0.99 mol
A⁻ → 1.00 mol + 0.010 mol = 1.01 mol

The pH of a 1.0 L solution of $c_a = 0.99 \text{ mol/1.10 L}, c_b = 1.01 \text{ mol/1.10 L}$ is ...

$c_a/c_b = 1.00 \rightarrow 0.99/1.01, \text{ pH} \rightarrow 5.01$ (tiny change!)

At 25 °C, the pH of a 1.0 L solution pure water, $K_c = 1 \times 10^{-14}$ is ...

pH = 7.00

Add 100. mL of 0.100 M NaOH ...

$[OH^-] = 0.010 \text{ mol/1.10 L} = 0.0091$

pOH = 2.04, pH = 11.96 (huge change!)
Different amounts of “not enough” base

At 25 °C, the pH of a 1.0 L solution of \( c_a = c_b = 1.00 \text{ M} \), \( K_a = 1 \times 10^{-5} \) is ...

\[ \text{pH} = 5.00 \]

Add 100 mL of 0.100 M HCl ...

\[ \text{HCl(aq)} + \text{HA}^-(aq) \rightarrow \text{H}^+(aq) + \text{Cl}^-(aq) \]
\[ \text{HA} \rightarrow 1.00 \text{ mol} + 0.010 \text{ mol} = 1.01 \text{ mol} \]
\[ \text{A}^- \rightarrow 1.00 \text{ mol} - 0.010 \text{ mol} = 0.99 \text{ mol} \]

The pH of a 1.0 L solution of \( c_a = 1.01 \text{ mol}/1.10 \text{ L}, c_b = 0.99 \text{ mol}/1.10 \text{ L} \) ...

\[ \frac{c_a}{c_b} = 1.00 \rightarrow 1.01/0.99, \text{ pH} \rightarrow 4.99 \text{ (tiny change!)} \]

Buffers

Mixtures of a weak acid and its conjugate base (or weak base and its conjugate acid) exhibit the special property that they resist changes to pH.

For this reason such mixtures are known as buffers.

Buffers resist change in pH

Added strong base (say, OH\(^-\)) is gobbled up ...

\[ \text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l) + \text{TA}^-(aq) \]
\[ c_a \text{ lowered, } c_b \text{ raised, } c_a/c_b \text{ lowered} \]

Added strong acid (say, HCl) is gobbled up ...

\[ \text{HCl(aq)} + \text{IA}^-(aq) \rightarrow \text{HA}(aq) + \text{Cl}^-(aq) \]
\[ c_a \text{ lowered, } c_b \text{ raised, } c_a/c_b \text{ raised} \]