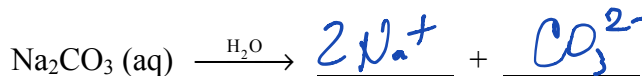
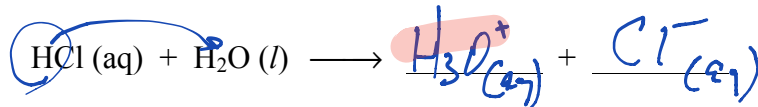


Rainbow Demo: Add saturated NaCO_3 (aq) to 0.1 M HCl (aq) containing Universal Indicator

Solutions:



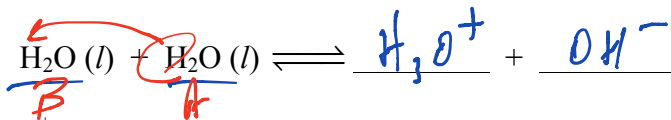
Rainbow Reactions:

RED ORANGE	<p>Excess Acid, Solution is Acidic (pH < 7):</p> $\text{HCl (aq)} + \text{H}_2\text{O (l)} \longrightarrow \text{H}_3\text{O}^+ + \text{Cl}^-$ <p>$[\text{H}_3\text{O}^+]$ is (high, low); $[\text{OH}^-]$ is (high, low)</p>
GREEN	<p>Neutral (pH = 7)</p> $\text{H}_3\text{O}^+ \text{ (aq)} + \text{CO}_3^{2-} \text{ (aq)} \rightleftharpoons \text{H}_2\text{O} + \text{HCO}_3^-$ <p>$[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ are equal</p>
BLUE VIOLET	<p>Excess Base, Solution is Basic (pH > 7):</p> $\text{CO}_3^{2-} \text{ (aq)} + \text{H}_2\text{O (l)} \rightleftharpoons \text{HCO}_3^- + \text{OH}^-$ <p>$[\text{H}_3\text{O}^+]$ is (high, low); $[\text{OH}^-]$ is (high, low)</p>

Thus, the relationship between $[\text{H}^+]$ and $[\text{OH}^-]$ is (direct, inverse).

Self-Ionization of Water

Even in neutral water, there is some $[\text{H}^+]$ and $[\text{OH}^-]$. Why?



Thus, in neutral water, $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ are explained by the following relationship:

$$K_w = 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-], \text{ so } [\text{H}_3\text{O}^+] = [\text{OH}^-] = \sqrt{1 \times 10^{-14}} = 1 \times 10^{-7}$$

So, since $1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$,

- As acid is added, $[\text{H}_3\text{O}^+]$ (increases, decreases) and $[\text{OH}^-]$ (increases, decreases)
- As base is added, $[\text{OH}^-]$ (increases, decreases) and $[\text{H}_3\text{O}^+]$ (increases, decreases)

If you know $[\text{H}^+]$ or $[\text{OH}^-]$, you can determine other: $[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{[\text{H}_3\text{O}^+]}$ and $[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]}$