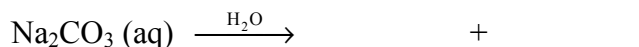
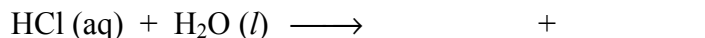


Rainbow Demo: Add saturated NaCO_3 (aq) to 0.1 M HCl (aq) containing Universal Indicator

Solutions:



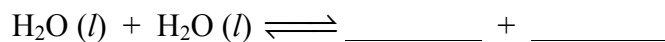
Rainbow Reactions:

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

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

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