

pH elimination: using a fixed amount of nutrients with certain pH and 1 L of water.

we have 5 ml of a solute, and we want to mix it with 1 L of water with a pH of 7.0, so that the resulting pH is approximately 6.0.

We can approach this problem using the concept of dilution and the relationship between pH and the concentration of hydrogen ions ($[H^+]$).

For dilute solutions, the pH can be approximated using the equation:

$$pH = -\log_{10}([H^+])$$

Initially, the pH of the water is 7.0, which corresponds to a $[H^+]$ of 10^{-7} moles/L. we want to mix in 5 ml of a solute, so the final volume becomes $1\text{ L} + 0.005\text{ L} = 1.005\text{ L}$.

Let's assume the solute has an initial pH of x . After mixing, the moles of $[H^+]$ from the solute and the water must balance out to achieve a pH of 6.0:

$$([H^+] \text{ from solute}) + ([H^+] \text{ from water}) = \text{Total } [H^+] \text{ after mixing}$$

Using the dilution equation ($C_1V_1 = C_2V_2$), where C is concentration and V is volume, the $[H^+]$ from the water remains the same (10^{-7}), and the $[H^+]$ from the solute is given by 10^{-x} (since the pH of the solute is x).

So the equation becomes:

$$10^{-x} + 10^{-7} = 10^{-6}$$

Solving for x :

$$10^{-x} = 10^{-6} - 10^{-7}$$

$$x = -\log_{10}(10^{-6} - 10^{-7})$$

$$x \approx 5.301$$

Therefore, the pH of the solute should be approximately 5.301 to achieve a resulting pH of around 6.0 after mixing it with 1 L of water with a pH of 7.0.