Buffers (pH changes by adding H^+)

Background

As shortly explained in the section about how to calculate pH in a solution of dissolved $NaHCO_3$ a buffer has the capability of keeping pH within a certain and narrow range even if an excess of hydrogen ions H^+ or hydroxide ions OH^- are added.

Before I move on I suggest you read the pdf documents showing how equilibriums are calculated and how pH in a $NaHCO_3$ solution is calculated. Links are found in the margin to the left.

Consider the sodium bicarbonate ($NaHCO_3$) solution

Consider the solution mentioned made of 0.1 mol/L $NaHCO_3$. We can show that unlike pure water this solution of sodium bicarbonate is extremely good at resisting pH changes when HCl is added. In fact, by adding 0.0005 mole of HCl it can be calculated, that the drop in pH in bicarbonate compared to pure water is negligible.

To follow the arguments an excel spreadsheet is needed to do all the calculations. You can download the spreadsheet at the www.phscale.net webpage.

As is seen in this excel spreadsheet at the end of row 16, the pH in the solution is 8.31. Now let's figure out what happens when 0.0005 mole HCl is added (volume change is neglected).

This is done by setting up a new charge balance that also includes $[Cl^-]$. By adding the HCl the [H+] is calculated from the equation:

$$[Na^+] + [H^+] - [Cl^-] - [OH^-] - \frac{[TIC]}{1 + \frac{[H^+]}{K_a^{CO_2^*}} + \frac{K_a^{HCO_3^-}}{[H^+]}} - 2 \cdot \frac{[TIC]}{1 + \frac{[H^+]^2}{K_a^{CO_2^*} \cdot K_a^{HCO_3^-}} + \frac{[H^+]}{K_a^{HCO_3^-}}} = 0.$$

The calculation for this equations starts at row 20 in the excel document. In row 29 it is seen that the pH should decrease to approximately pH = 8.21 which is not much.

Had it been a solution of pure water the pH would have dropped to $3.3 = -LOG_{10}(0.0005)$

A. Nielsen, March 21, 2008