by Bob Taylor

THE pH VALUE OF RAW WATER IS MEANINGLESS

The pH of water rarely indicates how much acid or alkali is needed to change the pH.

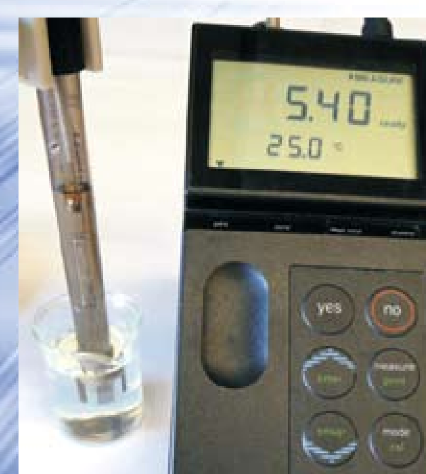
For example, it is not uncommon to have two different samples of water of equal pH where one requires four times more pH adjustment than the other.

This phenomenon is due to the concentrations of bicarbonate and carbon dioxide present in the water. It is particularly pronounced with bore waters.

Bicarbonate

Bicarbonate (HCO3-) is alkaline and, therefore, elevates pH. Its concentration is normally expressed as alkalinity. It is one of the main factors causing pH to rise in nutrient solutions and also confuses growers in their attempt to estimate how much pH adjustment solution will be required to lower pH.

Unlike hydroxide, bicarbonate is only weakly alkaline and, therefore, unable to elevate pH above 8.3, regardless of its concentration. As a consequence of this, unlike hydroxide, bicarbonate has a strong pH buffering capacity, which means it resists pH change when acid is added. For example, a weak solution of hydroxide can have a pH of 14.0 whereas a bicarbonate solution 10 times more concentrated has a pH lower than 8.3. Now, the interesting fact is, to lower the pH down to 4.5 the bicarbonate solution requires 10 times more acid than the hydroxide solution - even though its initial pH was so much lower.



The pH value of raw water is meaningless in most situations.

Hence, the presence of bicarbonate is deceiving because unlike hydroxide it is not detectable from pH readings and is only noticeable once you attempt to lower the pH.

High CO_2 levels are evident from the formation of bubbles in a bore water sample. The pH will rise as the CO_2 escapes

Carbon Dioxide

Have you ever wondered why pH fluctuates, typically upwards after it is lowered? This behavior is actually a consequence of adjusting the pH. Lowering pH via adding acid removes bicarbonate and produces carbon dioxide. The presence of this free, uncombined carbon dioxide (CO_2) tends to lower the pH because it reacts (only weakly) with water to form carbonic acid. However, CO_2 concentrations above about 0.5 milligrams per liter in water are unstable when such waters are exposed to the atmosphere (at sea level pressures). Under that condition, CO_2 in excess of 0.5 milligrams per liter will slowly escape from the water into the atmosphere. Consequently this loss of acidity causes a corresponding rise in pH.

This subsequent rise in pH is particularly noticeable with ground waters (i.e. bore water), which typically have CO_2 contents around 50 to 200 milligrams per liter (due to biological activity within the aquifer). When these waters are pumped to the surface, the pH rises with time because the excess (acidic) CO_2 gradually escapes (Fig above). The pH will then rise to a stable value solely dependent on the water's bicarbonate content.

Example: A bore water with 100 milligrams per liter bicarbonate and 100 milligrams per liter of free CO_2 will have an initial pH of 6.3. Its pH will gradually rise to 8.2 after it has been exposed to the atmosphere for sufficient time to allow the CO_2 content to drop to around 0.5 milligrams per liter.

The same phenomenon (although to a much lesser extent due to lower CO_2 contents) can occur with scheme (tap) water. Thus the conclusion - because the pH of waters is only stable after aeration, it is only the "after aeration" pH value that has any interpretative significance. To determine that value, aerate the water by tumbling a sample of it from one container to another, 30 to 40 times prior to measuring its pH.

Conclusion:

Interpret pH values with caution because water with a lower pH than another may produce the higher pH after both are aerated! MY

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