pH:CEC Ratio

A new approach to predicting agronomic potential

by Lawrence Mayhew

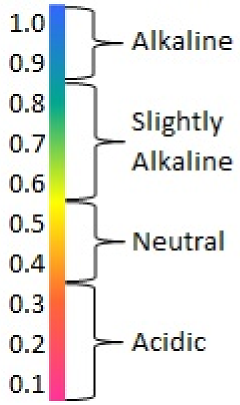
Soil pH adjustments and plant nutrient ratios based on cation exchange capacity have been used successfully for many decades to support more sustainable approaches to crop production. Recently, the mathematical relationship of soil pH, cation exchange capacity (CEC), and stable soil organic matter content of soils has been successfully applied to predict the potential for agronomic productivity for soils. The calculated results appear to have predictive value for crop production when bio-compatible inputs are used to correct mineral imbalances.

This review is an attempt to reveal the possible mechanisms that may be involved in the relationship of soil solution pH, CEC, and soil organic matter that provide the apparent predictability for potential agronomic production. CEC and pH are explored in the context of regenerative soil systems where the interaction of microbes, plant roots, clays, organic carbon substances, carbon dioxide, and *biomineralization* contribute to thermodynamically favorable energy transfers through soil systems. The importance of carbon dioxide released by plant roots and microbes, and microbial biomineralization of calcite are explored because these processes are all *charge balancing* mechanisms responsible for controlling nutrient ions adsorbed onto soil colloidal surfaces, buffering the underground soil environment making soils fit for life.

The Calculation and Unified Scale

The proposed algorithm for agronomic production potential (Ag_{p}) is applicable for soils that have 2% or more stable (dark) organic matter (OM). Because stable organic matter is derived from the breakdown of plant residues, fields that have undecomposed plant residues that are not more than 1 ½ years old are considered to be candidates for this calculation.

$$\mathrm{Ag}_{\mathrm{p}} = \sum \left\{\frac{pH}{CEC}\right\} + \left\{\frac{(OM-2)}{100}\right\}$$



Calculated values that fall in the range of 0.4 to 0.8 are predictive of optimal crop production, if stable soil organic matter content is equal to or greater than 2%. Ag_{p} results of less than 0.4 indicate acidity caused by a relatively high hydrogen (H^{+}) and aluminum (Al^{3+}) activity reducing the bioavailability of cation nutrients. The high acidity can be corrected by the application of "liming" agents, such as pulverized limestone rock. Bio-compatible inputs that do not disrupt pH and soil biological systems are applied to soils with Ag_{p} ratios between 0.4 and 1.0.

pH

The "p" in pH is the symbol for negative logarithm and the "H" is for hydrogen ion. pH is defined as the negative logarithm of the hydrogen ion *activity* in water solution;

$$pH = -log_{10}\left(a_{H^{+}}\right)$$

In the equation above, $a_{H^{+}}$ is the hydrogen ion activity. Notice that the word "activity" is used instead of *concentration*. Many people think that pH means the "potential" of hydrogen ions in solution; that's somewhat accurate. When soil pH is measured by a lab, it is done with an instrument that detects what is generally reported to be the hydrogen H^+ ion activity of the sample soil. A suitable definition for pH is a measurement of the *effective concentration* of hydrogen ion (H^+) in the solution, in other words how much of the H^+ manages to muscle its way through a very crowded environment of other cations and anions in solution. However, actual soil solutions contain many other chemicals besides hydrogen ions.

Soil pH measurements were invented to determine the amount of limestone needed to counteract the unfortunate acid-forming effects of over application of nitrogen fertilizers.[1] Although there is good correlation between conventional agronomic production and soil pH, the measurement alone has limited predictive value when used in the context of sustainable or regenerative agronomic production systems.

In natural soil systems, hydrogen ion concentration is part of the overall *charge balance* that must be regulated within a suitable pH range compatible with living systems. The maintenance of a suitable pH range is called *buffering*, which is accomplished by *biogenic* compounds; substances created by biological activity. Soil solution pH is naturally buffered by bicarbonates (HCO_3^-) and carbonates (CO_3^{2-}) released by plant roots (exudates) and microbiological interactions to *charge balance* soil acidity.

$$H^+ + CO_3^{2-} \leftrightarrow HCO_3^-$$

The origin of these buffering agents is actually carbon dioxide (CO_2) generated by microbes and plant roots, which instantly converts to carbonic acid when in contact with water.

$$CO_2 + H_2O \rightarrow H_2CO_3$$

In conventionally managed soils, where nitrogen fertilizers are typically used to supply nitrogen to plants instead of microbes, high rates of nitrogen inputs, such as urea, dissolve organic matter and generate large quantities of acid (H^+). The loss of organic matter and the increase in solution H^+ results in a depletion of exchangeable nutrient cations on the colloidal surfaces [2] in order to maintain a zero charge balance on the surfaces of the soil colloids. Additionally, high input rates of phosphate (PO_4^{2-}), sulfates (SO_4^{2-}) and chloride (Cl^-) anions can disturb the charge balance of soil solutions to such a degree that soils react to their presence by precipitating them out of soil solution in order to maintain chemical equilibrium.

Liming a soil with limestone ($CaCO_3$) to decrease acid conditions is illustrated by the chemical reaction below. Instead of microbes and plants supplying the carbonate anion (CO_3^{2-}) to balance acidity, limestone is used as a source of carbonate (CO_3^{2-}) to neutralize the hydrogen cations (H^+) responsible for acidity.

$$2\mathrm{H}^+ + \mathrm{CaCO}_3 \rightarrow \mathrm{Ca}^{2+} + \mathrm{H}_2\mathrm{O} + \mathrm{CO}_2$$

Chemical reactions that *buffer* soil solutions are reactions that create a state of *equilibrium*, meaning they restore chemical balance.

Humans use a number of different chemical buffers for their stomachs. For example, if a person takes sodium bicarbonate ($NaHCO_3$) for an acid stomach, they are taking advantage of the ability of the bicarbonate anion to react with the excessive H^+ cations responsible for the acid condition.

$Na^+ HCO_3^- + H^+ Cl^-$ ---> $Na^+ Cl^- + H_2O + CO_2$

The reaction between the hydrochloric acid ($H^+ Cl^-$) in the stomach and the sodium bicarbonate results in a sodium chloride (a salt), water, and carbon dioxide. When the reaction is complete, you belch; that’s the carbon dioxide.

However, it takes a lot of energy to break the Ca-CO_3 chemical bonds of limestone; that’s why the reaction goes forward very slow and over long time periods after application to soil. Additionally,, the efficiency of the reaction decreases as the presence of the strong acid decreases over time. The reaction appears to provide an “available” soluble form of calcium (Ca^{2+}) for plant consumption, but in soils with less than 3% organic matter the calcium (Ca^{2+}) is free to react with other soil anions in the geological environment, such as phosphates (PO_4^{2-}), to form chemically stable compounds that precipitate out of soil solution. In natural soil systems, sufficient with humic matter and microbial activity, calcium is kept in a more bioavailable form, preventing precipitation out of soil solution.[3]

Measuring Soil pH

When laboratories measure the pH of a soil, the soil sample is diluted with pure water, not actual soil solution water. As natural soil solutions are populated with innumerable positively charged cations and negatively charged anions that compete with hydrogen (H^+) protons [7] for space on colloidal clays and organic matter, especially humic substances[4], there is a significant difference between hydrogen ion (H^+) activity in a test tube filled with pure water and the activity of natural soil solutions. Using the CEC illustration in Figure 1 as an example, H^+ (black spheres) and four additional cations are shown in the solution; in a real soil solution there are hundreds of possible combinations of cations [5] capable of exchanging on colloidal surfaces competing with H^+ protons for a position on the colloids. Calcium (Ca^{2+}),magnesium (Mg^{2+}), sodium (Na^+), and potassium (K^+) cations and chlorides (Cl^-), sulfates (SO_4^{2-}), carbonates (CO_3^{2-}), fluorides (F^-) anions compete *electrostatically* with each other in natural soil solutions. This is called *steric hindrance;* also called *crowding effects.*

The concept of steric hindrance, crowding effects, limiting the activity of the H^+ hydrogen ions is analogous to a person entering a room that contains other people. If the number of people in the room is very small, the person is free to move about in any direction, making their way to places where they can perform their "activities". However, if the room is very crowded, the person has a difficult time getting to the bathroom, for example. This is analogous to the movement that limits the activity of ions in soil solution.

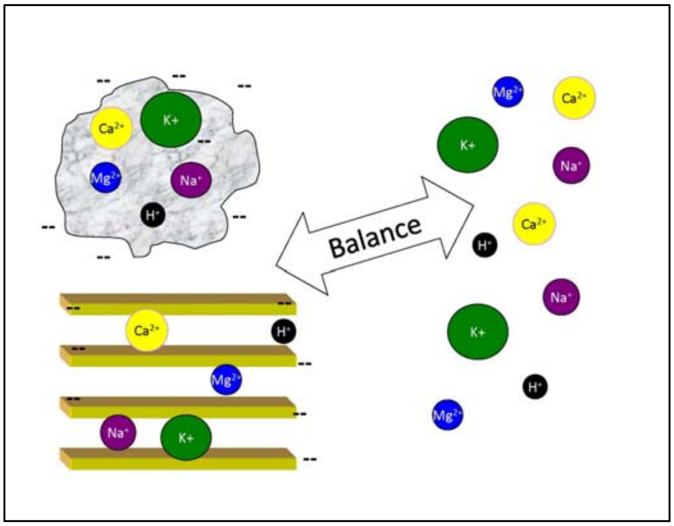
Soil solutions are a combination of dissolved chemicals and solid *suspensions* of colloidal materials that are not soluble in water. The presence of colloidal clays and humic substances also buffer soil pH. The potential for these colloids to buffer soil solutions prepared for laboratory conditions is well known. The practice of allowing a mixture of soil and water to settle out solid particles before measuring pH and making sure the probe is only in the clear solution above the sediment was first noted 1931.[6] This is another "ideal" condition that does not exist in soil solutions.

* * *

Another common soil cation that competes with H^+ for adsorption surfaces is aluminum. Aluminum cations (Al^{3+}) are capable of displacing other cations on the absorption surfaces of colloids as well as H^+ protons.[1] When measuring pH in a lab, the pH measurement does not compensate for the interference created by Al^{3+} activity, therefore aluminum (Al^{3+}) cations are included in the pH measurement. However, in natural soils systems, humic substances reduce the activity of aluminum. Therefore, the pH:CEC ratio is adjusted by the presence of stable (humic) organic matter.

CEC

Figure 1. Exchange of cations in soil solution with organic matter and clays.



In the illustration (Figure 1.) imagine that the clear white space around the objects represents the *soil solution* where calcium (Ca^{2+}), magnesium (Mg^{2+}), sodium (Na^+), potassium (K^+), and hydrogen (H^+) ions are exchanged between the soil colloids and the water. The negatively charged grey colored blob on the upper left represents colloidal humic substances, and the layered negatively charged plates on the lower left are colloidal clays. The cations on the right side of the page are dissolved in the soil solution, making them "available" nutrients for both microbes and plant roots.

Cation Exchange Capacity (CEC) measures the amount of positively (+) charged plant nutrient cations that can be absorbed on to the surface of soil clays and humic substances. Positively charged hydrogen ions (H^+) can adsorb to these surfaces also, but the chemical activity of the H^+ ions is significantly different than the activity of nutrient cations, such as Ca^{2+} or Mg^{2+}, having a profound impact on the chemistry of soil solutions and agronomic production.

The key concept of CEC is positively charged cations in soil solution are exchanging places with loosely bound (*adsorbed)* cations on surfaces of clays and humic substances, making the cations available as nutrients for both microbes and plant roots that are in contact with the soil solution. The CEC is usually dominated by calcium (Ca^{2+}), magnesium (Mg^{2+}), sodium (Na^+), and potassium (K^+). However, when soils are too acid, that is an indication that there are potentially toxic levels of aluminum cations (Al^{3+}) as well as too many hydrogen protons (H^+) adsorbed on to soil colloids and in soil solution.

Notice that the number of cations adsorbed onto the humic substances and the clays are equal to the number of cations in the soil solution; they are balanced. If, for example, either a microbe or a plant root removes any cations from the soil solution, they will have to be replaced by an equal number of cations to maintain chemical balance (equilibrium) between the soil solution and colloids. The cations that come from the surface of the colloids are regarded as *exchangeable* cations, i.e. loosely held on the charged surface of the colloids for the purpose of being easily released into the soil solution where they are consumed by living things.

The number and amount of charge on the cations are essentially maintaining a net zero charge across the colloidal surfaces. The need to maintain a zero point of charge does not mean that the surface has no charge, but rather there are equal amounts of positive and negative charge on the surface of the colloids. In natural soil systems, cations are replenished by the action of plant root exudates and microbial acids through the breakdown of insoluble minerals. In other words, the charge balance of soil systems is under biological control.

Measuring Soil CEC

The measurement of CEC is closely entwined with the measurement of pH. The two measurements are interrelated because there is little difference between the assumptions behind their *operational* definitions. The capacity of soils to hold positively charged cations originates from the negative charge carried on soil particles, primarily from clays and the dark organic matter called humic substances. As this phenomena cannot be observed directly with the human eye, the definition for measuring what is perceived to be the total amount of positively charged cations that are capable of adsorbing onto the surfaces of soil particles is defined *operationally,* meaning it is defined by the operation by which it is measured, which introduces many variables and assumptions regarding the mechanisms.

There is no standardized method for measuring soil CEC. In fact, there are numerous analytical procedures for calculating CEC, each one designed to overcome the effects that seem to have an impact on the outcome of the analysis.

However, both pH and CEC are measurements of cations; pH measures the *activity* of H^+ ions in solution, while CEC measures the *amount* of cations that may be adsorbed on soil surfaces. Activity is a thermodynamic measure of the "effective concentration" in a water mixture. Effective concentration is the amount of a chemical needed to overcome all the chemical interactions in water solution to have a desired effect. Many of the conditions that affect the outcome of pH measurements also impact the measurement of CEC, especially the "crowding effects".

pH to CEC Relationships

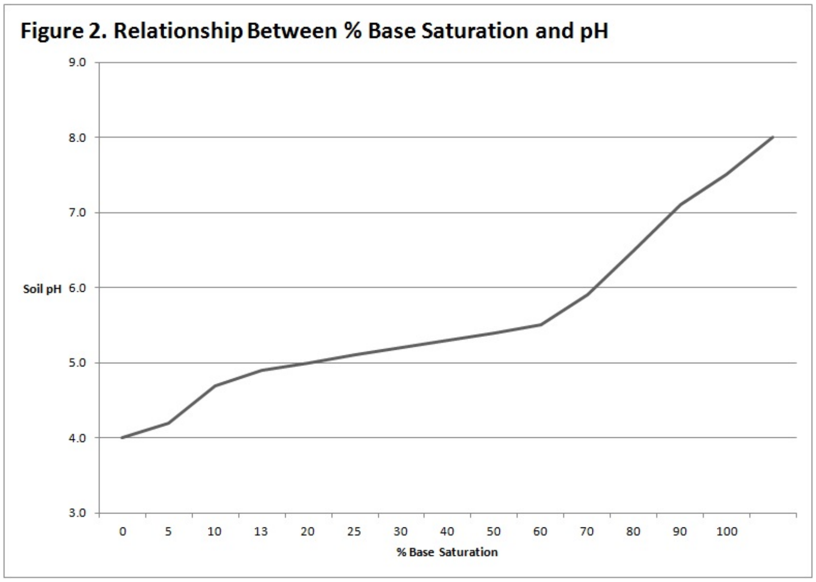


Figure 2. Relationship Between % Base Saturation and pH

The relationship of pH to CEC has been extensively studied. In one study[7], 200 different soils were tested for pH and CEC. The results represented in Figure 2 demonstrate a linear relationship between pH and CEC when pH was between ph 4.5 and 6, with only a small variation in that relationship outside of that pH range. This linear relationship is easily extrapolated into a series of ratios of pH:CEC.

Early studies on the relationship of ph to CEC were conducted to determine how much limestone (calcium carbonate) was needed to neutralize acid soils, as the CEC of a soil will affect how much calcium (Ca^{2+}) can be adsorbed by soil colloids. Early researchers assumed that the amount of Ca^{2+} released by the reaction of limestone ($CaCO_3$) with acid (H^+) produced sufficient "available" calcium in soils.

$$2H^+ + CaCO_3 \rightarrow Ca^{2+} + H_2O + CO_2$$

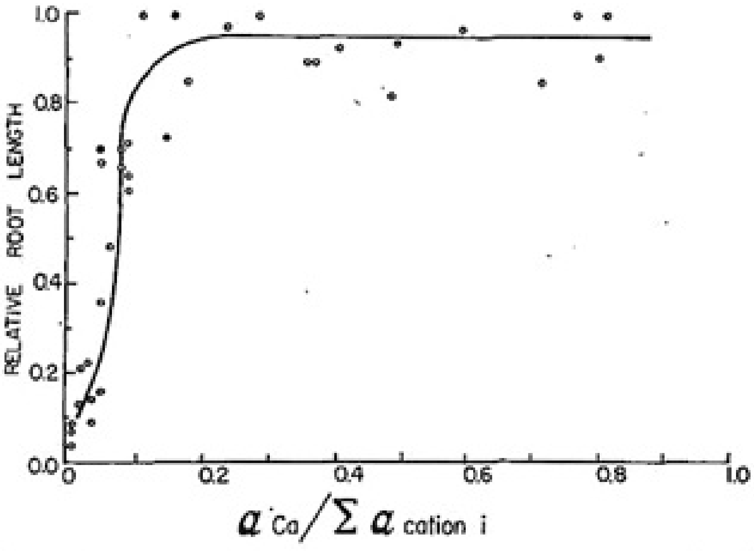
However, there are many instances of calcium deficiency in limestone treated soils. The term “deficiency” used by early researchers[8] is not the same as the concept of balancing nutrients for optimal effect used in regenerative agriculture. Conventionally, yield is the primary measurement used to determine if there is sufficient calcium in a soil. However, from a systems viewpoint, calcium needs to be supplied as part of an overall nutrient balance plan where greater root development leads to better nutrient access, preventing lodging, and encouraging sufficient root exudates to the soil microbial population to release more nutrients. Optimal amounts of calcium (Ca^{2+}) also provide protection from diseases and insects.

In 1966, Professor Fred Adams conducted an experiment to determine why cotton seedlings fertilized with ammonium nitrate were not expressing proper root development.[9] He observed that the root development was not related to soil pH, or the rate of phosphorus and ammonium applied, nor absolute concentrations of cations in the soil solution. When Professor Adams compared the ratio of the activity of the total cations to the activity of the calcium cations ions in soil solution with root growth, the relationships became evident. Figure 3 shows how the ratio of calcium (Ca^{2+}) cation activity relative to the activity of the total cations in the soil solution affects root growth. The same relationship was revealed when the ratio of ammonium (NH_4^+) cations to total cations in the soil solution was compared to root growth.

The amount of calcium in a soil available to plants is closely tied to a soil’s geology, organic matter, and carbonic acid released by microbes, more correctly referred to as *geobioavailable* [a]calcium. Geobioavailable calcium is not necessarily provided by the application of limestone rock because it originated as a non-crystalline form of calcium carbonate that has undergone the rock forming processes of compaction.

Wolt, D.W., 1994. *Soil Solution Chemistry, Applications to Environmental Science and Agriculture.* John Wiley & Sons, New York, pp. 217-219.

Fig. 3. Effect of calcium: total cation activity on initial growth of primary roots of cotton seedlings. Adams 1966.



The Adams (1966) experiment was repeated four years later using sudangrass and cotton. In both experiments, the data revealed;

- the activity of soil solution calcium is reduced by its interaction with other cations that antagonize calcium cations (crowding effect);

- the ratio of calcium to other cations is more important than concentration of calcium alone;
- calcium cation activity is not raised by "liming" soils;
- calcium deficiencies become evident when they are expressed as suppression of root development instead of yield;
- ratios of calcium to total cation activity above 0.2 increased relative root growth;
- ratios of calcium to total cation activity below 0.15 resulted in root death.

The relationship of calcium cation activity relative to total cation activity revealed by Adams (1966) is strikingly similar to the proposed pH:CEC algorithm and the results from Adams (1966) are similar to the proposed unified scale of this paper.

Adams (1966) expressed the relationship of calcium to CEC as the following algorithm;

$$a_{\text{Ca}} / \sum a \text{ cation i}$$

The proposed algorithm that is the subject of this paper is;

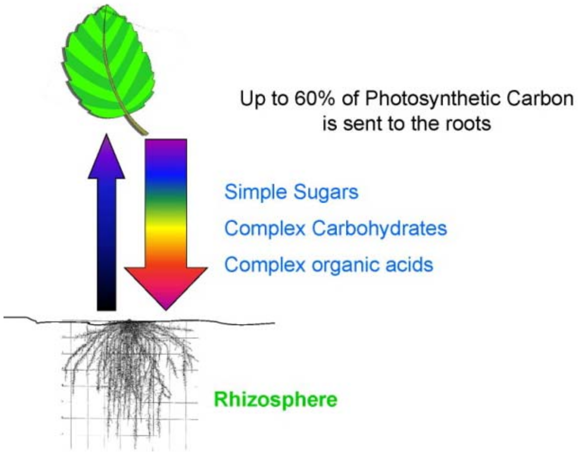
$$\text{Agp} = \sum \left\{\frac{pH}{CEC}\right\} + \left\{\frac{(OM-2)}{100}\right\}$$

Both calculations compare the activity of one species of positively charged cations in soil solution to the total of exchangeable cations. The proposed algorithm in this paper uses hydrogen ion activity (pH) in place of calcium ion activity (a_{Ca}), and the sum total of exchangeable cations ($\sum a_{\text{cation i}}$) is expressed as the CEC; essentially the same.

Organic matter (OM) in the form of humic substances influences soil pH, helping to keep soil pH within certain boundaries; the effect is called buffering. The distinguishing characteristic between the proposed and the Adams algorithms is the adjustment of the pH:CEC ratio by stable organic matter.

One possible aspect of pH:CEC ratio is that it may be an indicator of the *bioavailability* of cation nutrients in soil solution because of the direct relationship between pH and CEC 7 and the bioavailability of nutrients expressed as a ratio of total cations to hydrogen ions in solution.9 It is also well known that the CEC of a soil has a certain capacity for H^+ and aluminum cations (Al^{3+}) as well as nutrient metals, and humic substances have a preference for removing potentially toxic aluminum (Al^{3+}) cations out of soil solution. All of these factors are closely related to the charge balance mechanisms of soils, a relationship that seems to be unified by the pH:CEC algorithm.

Biological Buffering



Soils are an incredibly complex mélange of chemical, biological, and physical interrelationships that have evolved over extremely long periods of time into elegant self-balancing systems that provide balanced nutrition for plants, animals, and humans. Long before there were plants or humans, microorganisms saturated the below-ground atmosphere with carbon dioxide as well as oxygenating the above-ground atmosphere. This set the stage for geomicrobiological interactions that break down rocks into minerals (weathering) that are then broken down further into nutrients used by both plants and soil microbes.

Plants contribute to this system by capturing carbon dioxide (CO_2) and water (H_2O) from the above-ground atmosphere converting it into an endless variety of complex organic carbon compounds that are sent to their root system and released as root exudates that breakdown minerals, feed soil microbes, and contribute to the buildup of soil organic matter. Plant roots also release negatively charged bicarbonate anions $(HCO_3)^-$ and CO_2 that are critical to *buffering* the soil system pH.

In natural systems, dying plants are recycled into the below-ground system where they are broken down by an extensive array of microbes until the only substances left are complex carbon compounds that recombine with themselves into the most complex carbon substances known as humic substances. This process is called *humification*; the products of humification are incredibly complex carbon-based (organic) substances that are the primary constituents of the dark colored matter of *humus.* They are extremely stable (recalcitrant) substances that resist any further breakdown by microbes, potentially lasting for thousands of years. They provide the conditions necessary for many life-giving geomicrobiological soil interactions to occur primarily by holding water, detoxifying harmful substances, and providing negatively charged (-) *colloidal* surfaces for the exchange of positively charged (+) nutrients (cations) with other cations in the *soil solution*. For the purpose of this paper, colloids are defined as insoluble soil particles ≤2μm in size that are suspended in or in contact with the soil water solution.

Humic substances are composed primarily of water-insoluble materials that operate in close proximity to clays, which are water-insoluble also. These two nearly inseparable materials interact with each other as the conductors of an underground symphony of life. The potential of humic substances and clays to provide the right conditions to release nutrients for microbes and plants is measurable; the measurement is called *cation exchange capacity* (CEC), which along with pH, are the most studied aspects of soil systems. Both humic substances and clays are powerful buffers.

Carbon dioxide

Ironically, the single most important inorganic carbon compound controlling all the great biological cycles that occur underground is rarely characterized that way. It is carbon dioxide (CO_2). The combination of water and carbon dioxide, called carbonic acid (H_2CO_3), is the basis of the limitless number of all life giving organic compounds. Along with calcite, a form of calcium carbonate, they are responsible for buffering all the global soils, rivers and oceans. Calcite and carbonic acid are the most important factors that make the environment fit for the purpose of life.[10]

You are probably familiar with the concept of photosynthesis, which is the process of plants capturing CO_2 from the atmosphere and combining it with H_2O in the presence of sunlight and chlorophyll to produce sugars. To illustrate the concept, the chemical reaction may be illustrated as;

$$CO_2 + H_2O \xrightarrow{\textit{Chlorophyll}} C_6\,H_{12}\,O_6$$

Some of the CO_2 captured by plants is photosynthesized into extraordinarily complex carbon compounds that are sent down to the plant root rhyzosphere as plant root exudates where they are released to attract beneficial microbes. Beneficial microbes have a dual role; repelling harmful soil microorganisms, and converting insoluble minerals into plant usable nutrients.[11] Of all the carbon captured by plant leaves, much of it is released as CO_2 into the soil by plant roots.[12] Upon contact with water, carbon dioxide rapidly changes to carbonic acid.

$$CO_2 + H_2O \rightarrow H_2CO_3$$

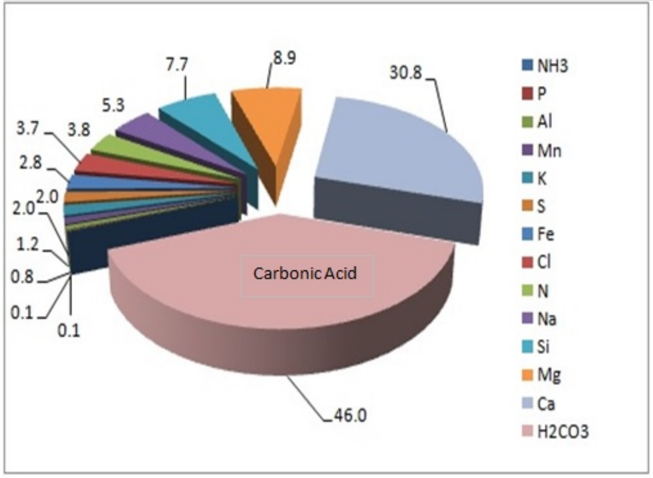
Carbonic acid easily dissociates in water into H^+ protons (acid) and bicarbonate anions (base). The log K -6.55 noted below next to the equation tells a chemist how the acid and base balance each other. The -6.55 indicates that carbonic acid dissociates at close to 7 pH, making it an effective buffering agent.

$$H_2CO_3 \rightarrow H^+ + HCO_3^- \qquad \log K\ -6.55$$

The combination of carbon dioxide and water as carbonic acid has the extraordinary capacity to prevent wide variations of pH (buffering) in all aqueous environments. Although CO_2 is reported to be over 100 times greater concentration in the soil atmosphere compared to the earth's atmosphere, there is sparse current data on how the concentration of CO_2 in typical soil solutions regulates the underground environment.

Going back into the literature to 1913, L. J. Henderson [13] estimated the average composition of rivers as having 46 parts per million concentration of CO_2 (Figure 3). Notice that the primary components of river water reported by Henderson are carbonic acid and calcium. Roughly, the same proportions of components appear in soil solution. In well drained soils with sufficient rainfall, carbonic acid is the predominant buffering agent,[14] buffering those soils within a pH of 5.5 to 6.5.[15]

Fig. 3. The Composition of River Water.
Modified from L. Henderson, 1913



As plant roots release acid as H^+ and simultaneously release bicarbonate (HCO_3^-) through root respiration, and as microbes are fed by plant root exudates, who in turn also releasing CO_2, it becomes apparent that as the size of a plant root mass increases, they have a major impact on soil pH buffering. As more H^+ are donated to the exchange sites responsible for CEC, more nutrient cations such as calcium and magnesium are released. This is because H^+ protons are preferentially adsorbed to cation exchange sites.1 And, bicarbonate anions participate in regulating the amount of cations released from cation exchange sites. Again, the combinations of all of these chemicals in the soil solution are considered to be *charge balancing mechanism*s that are under biological control.

Increased root mass allows more organic acids to be released by roots, such as carboxylic acids that are capable of buffing soil solutions in lower pH ranges of 3-5. The carboxylic acids are rapidly consumed by microbes as sources of carbon energy, releasing more CO_2, which in turn produces more carbonic acid to buffer the pH upward into a pH range that is more suitable to life. While microbes are dining on these energy rich carbon based acids, they pull some of the free calcium (Ca^{2+}) out of soil solution, combine it with carbonate (CO_3^{2-}) from the bicarbonates (HCO_3^-) in soil solution making a special form of calcium carbonate ($CaCO_3$).

Biomineralization

Many single celled organisms, bacteria, fungi, plant and animals engage in what is called biologically *controlled mineralization*.[16] The chemical pathways that start out using carbon dioxide and bicarbonates as the raw materials are under the direct control of microorganisms that direct the crystallization process in order to produce calcite [17] of a particular form whose crystalline orientation is done in such a way as to accommodate natural organic compounds and microbial adhesion; referred to as *chiral* crystals.[18] This particular form of calcite is made by soil microbes so that its chiral surface is recognized by microbes, allowing natural amino acids (that are also chiral) to adhere to its surface, which is part of the mineral cycling in soils, where the interactions with chiral biominerals is thermodynamically favored.[19]

The overall process of soil microbes intentionally producing a specific crystalline form of mineral fit for a specific purpose is called *biomineralization,* which is very common. Biomineralized chiral calcite is the most common form of calcium carbonate in soils. It is produced by soil microbes for the specific purposes of storing a bioavailable form of calcium and as the raw material for buffering soil solutions. Biomineralized products range anywhere from minerals intended to be consumed by microorganisms to those that are not intended to be consumed. The form of calcium carbonate that provides high mechanical strength, such as shells, teeth and bones is not crystallized. Non-crystalline forms of calcium carbonate that have no particular structure, such as the calcium carbonate of clam shells and exoskeletons of crabs and lobsters, are called *amorphous calcium carbonate*,[20, 16] produced primarily for mechanical strength and resistance to breakdown by microbes; that's why you can pick seashells washed up on to the seashore.

Bio-Compatible Materials

The predictability of the pH: CEC algorithm is effective when bio-compatible soil amendments are used. However, if a soil's pH is too acid, and if correcting the pH brings the ratio into the acceptable range, then limestone rock can be used to correct soil acidity. Otherwise, materials that are considered *bio-compatible* are much more effective.

Bio-compatible inputs are defined as chemically stable non-synthetic materials that are nearly insoluble in water, have surface structure and geochemical properties that are relevant to soil microbial interactions, and compatibility with the natural thermodynamics of soil systems that preserve the structure of water molecules and *hydrophobic* mechanisms[21], all of which participate in the natural carbon cycles of the globe.

Conclusion

The principles of nutrient balancing are based on the charge balancing relationships of soil nutrient cations, hydrogen protons (H^+), plant root exudates, geomicrobiological interactions, and stable soil organic matter. As the pH:CEC algorithm has been used successfully by practitioners for the last four years, it appears to be an advancement in the application of nutrient balancing practices used in regenerative agricultural systems, where the ratio of exchangeable soil solution components is more critical to crop production and protection than overall concentration of nutrients.

References

---

[1] Bloom, P.R., Skyllberg, U.L., and Sumner, M., 2005. Soil Acidity. In: M.A. Tabatabai and D.L. Sparks (eds.), *Chemical Processes in Soils.* Soil Science Society of America Book Series No. 8, Soil Science Society of America, Madison, Wisconsin, pp. 411-459.

[2] Barak et al., 1997. Effects of long-term soil acidification due to nitrogen fertile inputs in Wisconsin. Plant Soil 197:61-69.

[3] Stotzky, 1983. Influence of soil mineral colloids on metabolic processes, growth, adhesion, and ecology of microbes and viruses. In: P.M. Huang and M. Schnitzer (eds.), *Interactions of Soil Minerals with Natural Organics and Microbes.* SSSA Special Publication No. 17, Soil Science Society of America, Madison, Wisconsin pp. 305-428.

[4] Wolt, D.W., 1994. *Soil Solution Chemistry, Applications to Environmental Science and Agriculture.* John Wiley & Sons, New York.

[5] Essington, M.E., 2004. *Soil and Water Chemistry: An Integrated Approach*, CRC Press, Boca Raton.

[6] Wiegner, G., 1931. Some physicochemical properties of clay:II. Hydrogen clay. Journal of Society of Chemical Industry 50:103T.

[7] Peech, M., 1965. Lime Requirement. In: C.A. Black (ed.), *Methods of Soil Analysis, Part 2, Chemical and Microbiological Properties,* American Society of Agronomy, Madison, Wisconsin, pp. 927-932.

[8] Arnon, D.I. and P.R. Stout, 1939. The essentiality of certain elements in minute quantity for plants with special reference to copper. Plant Physiology 14:371-375.

[9] Adams, F., 1966. Calcium Deficiency as a Causal Agent of Ammonium Phosphate Injury to Cotton Seedlings. Soil Science Society of America Proceedings 30:485-488.

[10] Henderson, L. J., 1913. *The Fitness of the Environment: An Inquiry into the Biological Significance of the Properties of Matter.* The Macmillan Company, New York.

[11] Walker, T.S., et al, 2003. Root Exudation and Rhizosphere Biology. *Plant Physiology* 132(1): 44-51.

[12] Lambers, H., et al., 2002. Respiratory patterns in roots in relation to their functioning. In: Y. Waisel, A. Eshel and U. Kafkafi (eds.), *Plant Roots. The Hidden Half,* 3rd ed., Marcel Dekker, New York, pp. 521-552.

---

[13] Henderson, L. J., 1913. *The Fitness of the Environment: An Inquiry into the Biological Significance of the Properties of Matter.* The Macmillan Company, New York.

[14] Sposito, G., 2008. *The Chemistry of Soils,* 2nd Edition*.* Oxford University Press, New York.

[15] Bloom, P.R., Skyllberg, U.L., and Sumner, M., 2005. Soil Acidity. In: M.A. Tabatabai and D.L. Sparks (eds.), *Chemical Processes in Soils.* Soil Science Society of America Book Series No. 8, Soil Science Society of America, Madison, Wisconsin, pp. 411-459.

[16] Lowenstam, H.A. and S. Weiner, 1989. *On Biomineralization.* Oxford University Press, Oxford, England.

[17] Nielsen, M., S. Aloni, and J.J. DeYoreo, 2014. In situ TEM imaging of $CaCO_3$ nucleation reveals coexistence of direct and indirect pathways. *Science* 345(6201):1158-1162.

[18] Hazen, R.M., 2004. Chiral crystal faces of common rock-forming minerals. In: *Progress in Biological Chirality*, G.Palyi, C.Zucchi & L.Caglioti [Eds], Oxford, Elsevier, 2004. pp.137-151.

[19] Ryan, R., Myung, S. and Clemmer, D.E., 2005. Do homochiral aggregates have an entropic advantage? *Journal of Physical Chemistry B* 109:440-444.

[20] Addadi, L., Raz, S., and Weiner, S., 2003. Taking advantage of disorder: amorphous calcium carbonate and its roles in biomineralization. *Advanced Materials* 15(12):959-970.

[21] Tanford, Charles, 1997. How protein chemists learned about the hydrophobic factor. *Protein Science* 6:1358-1366.