

Brewing Water - An Overview

1.0 Introduction

This article is based on a presentation given by the author Craft Brewers Conference April 11 of last year in Cleveland. Space constraints here make it impossible to touch on all the topics discussed in Cleveland so the decision was made to concentrate on the material which is, in the author's opinion, at the heart of brewing water chemistry and that is the relationship between alkalinity (bicarbonate content), hardness (calcium and magnesium content), malt phosphate and mash pH. The importance of this subject stems from the importance of establishment of proper pH in the mash tun order to insure that it will be in the proper range during subsequent steps of wort and beer processing.

We will discuss the usual way in which bicarbonate (and hardness) get into brewing water, means by which brewers traditionally remove some of it and the mechanism by which malt phosphate "neutralizes" part of what remains. We will show how to estimate (from a chart) the quantity which is not so neutralized (Kohlbach's Residual Alkalinity) and its effect on mash pH.

While some of what follows may look frighteningly complicated it is in fact appreciably simplified. The intent is to convey an appreciation for the fundamental processes to the level where those who so desire can do some simple calculations to gain further insight. Accordingly some mathematical formulas are included without much explanation. These can be plugged into a simple spreadsheet (the author, at ajdel@cox.net, will be happy to e-mail an Excel example to interested readers). Those intimidated by the chemistry are encouraged to skip forward to the end of the article where the Residual Alkalinity chart is to be found. The sidebar has definitions of some of the chemistry terms used here.

Readers seeking additional information on brewing water can consult (Ref. 1.) and, for a more modern perspective, Ref 2.

2.0 Carbonic Acid, Bicarbonate Ion, Carbonate Ion

Establishment of strike pH depends upon the chemistries of two acids, carbonic and phosphoric and their largely insoluble calcium salts. Carbonic acid, with chemical formula H_2CO_3 contains two hydrogen atoms. A hydrogen atom is made up of two charged subatomic particles: a proton (positively charged) and an electron (negatively charged). The defining property of an acid (in this context) is that it can release the protons of its hydrogen atoms while retaining the electrons. The protons, are ionized hydrogen atoms symbolized by H^+ . The terms *hydrogen ion* and *proton* are used interchangeably in this article.

Carbonic acid gives up a proton in the reaction.



The electron which stays with the carbonic molecule lends it a net negative charge of 1 (indicated by the “-” sign). The carbonic molecule which has lost one proton is called a *bicarbonate ion*.

The double headed arrow in the chemical “equation” indicates that carbonic acid can shed a proton and become a bicarbonate ion or that a bicarbonate ion can pick up a proton and become a carbonic acid molecule. In other words, the reaction can take place in either direction and it does. All three species (carbonic acid, hydrogen ion and bicarbonate ion) are present simultaneously and their relative concentrations at equilibrium are governed by the law of mass action which, for this case, is

$$\frac{[H^+][HCO_3^-]}{[H_2CO_3]} = K_1 = 10^{-6.35} \quad (2.2)$$

The brackets indicate the concentrations of the components in moles per liter. K_1 is called the *equilibrium constant* and is related to the energy which binds the proton to the molecule.

In chemistry calculations are simplified appreciably by working with the negative logarithm of concentrations. The negative log is indicated by the use of the symbol p placed in front of the concentration. Taking the negative logarithm of both sides of Equation (2.2) gives

$$p[H^+] + p\frac{[HCO_3^-]}{[H_2CO_3]} = pK_1 = 6.35 \quad (2.3)$$

In this equation $p[H^+]$ means just what you probably suspect i.e. the pH . It's handy to label the ratio of bicarbonate concentration to carbonate concentration r_1 (with the one indicating that it applies to the reaction in which the first proton is lost). Doing this and rearranging we can write

$$pr_1 = pK_1 - pH \quad (2.4)$$

This is the most important relationship in brewing water chemistry. Appreciate it fully and the rest falls into place. To find the numerical value of the concentration ratio we use the inverse p operation which involves changing the sign and then taking 10 to the resulting power. Doing this to Equation (2.4) gives

$$r_1 = 10^{(pH - pK_1)} \quad (2.5)$$

This shows that if the pH is equal to pK_1 the ratio is equal to 1; that is that the concentrations of bicarbonate and carbonic are equal. If the pH is one unit larger than pK_1 the ratio will be 10 and the concentration of bicarbonate will be 10 times that of carbonic. If the pH is one unit less than pK_1 (i.e. about 5.35 within the range of desirable mash pH) then the ratio will be 0.1 meaning that the carbonic concentration will be 10 times the bicarbonate.

The bicarbonate still has a proton to give up. The reaction is:



The carbonic molecule from which both protons have been removed has a double negative charge and is called a *carbonate ion*. There is an equilibrium condition for this reaction as well

$$pH + p \frac{[CO_3^{2-}]}{[HCO_3^-]} = pK_2 = 10.38 \quad (2.7)$$

and the ratio of carbonate ion concentration to bicarbonate ion concentration, $r_2 = \frac{[CO_3^{2-}]}{[HCO_3^-]}$, can be calculated in exactly the same way as for the bicarbonate/carbonic ratio except that the value of the constant is different

$$r_2 = 10^{(pH - pK_2)} \quad (2.8)$$

with $pK_2 = 10.38$ at room temperature.

Given that we have the ratio of bicarbonate to carbonic, r_1 , and the ratio of carbonate to bicarbonate, r_2 , it is easy to compute the fraction of total carbo (*carbonic*, *bicarbonate*, *carbonate*) moles in a solution which are, respectively, carbonic, bicarbonate and carbonate. First compute $d = 1 + r_1 + r_1 r_2$. Then the fraction of moles of carbonic is $f_1 = 1/d$, the fraction which is bicarbonate is $f_2 = f_1 r_1$ and the fraction which is carbonate is $f_3 = f_2 r_2$. Readers may wish to calculate values for f_1 , f_2 , and f_3 for various values of pH and plot them.

3.0 How Bicarbonate and Calcium Get into Water

At this point we are ready to discuss how carbo species get into brewing water get into water in the first place. In most waters the source is dissolved calcium carbonate, $CaCO_3$, usually limestone. We mentioned earlier that calcium carbonate is not very soluble. When it is exposed to water the reaction



takes place. The underscore indicates that the species so labeled is a solid which is not in solution. This reaction also has an equilibrium constant associated with it. Since the $\underline{CaCO_3}$ is a solid its concentration can't vary and this fixed concentration is lumped into the equilibrium constant. The equilibrium condition is, therefore, written:

$$p[Ca^{+2}] + p[CO_3^{2-}] = pK_s = 8.45 \quad (3.2)$$

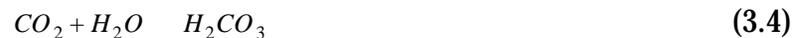
where the subscript s indicates that the constant applies to the solubility condition and the value 8.45 applies to room temperature conditions. Equation (3.2) actually describes the condition in which the solution is just saturated with calcium carbonate and represents the onset of precipitation. It is often cast in the form of the "solubility product".

$$[Ca^{+2}][CO_3^{2-}] \leq 10^{-pK_s} \quad (3.3)$$

If the concentrations of calcium and carbonate are such that this inequality is satisfied, precipitation will not occur.

In the simplest case initially pure water comes into contact with limestone so that each dissolved molecule puts one calcium ion and one carbonate ion into solution. Then the concentrations of carbonate and calcium will be equal and $p[Ca^{+2}] = p[CO_3^{2-}] = \frac{pK_s}{2} = 4.225$. The value of the carbonate concentration is thus $10^{-4.225} = 0.000059$ moles per liter or 3.5 mg/L, which isn't much. If hydrogen ions are available they will combine with the carbonate ion to become bicarbonate ion (Equation (2.6)) which allows more limestone to dissolve to replace the carbonate that converted. Equation (3.2) must still be satisfied. Note that though dissolved carbonate is "removed" by being converted to bicarbonate the same is not true for the calcium. If, for example, the process we are considering has proceeded to the point where a millimole of limestone has dissolved, then the calcium level will be 1 millimole or 40 mg/L. In this case the maximum carbonate concentration is limited to $p[CO_3^{2-}] = 8.45 - 3.00 = 5.45$ - more than an order of magnitude less than when we began with deionized water. But if the pH is low enough the ratio carbonate to bicarbonate is quite small. At pH 7.38 it is .001 so bicarbonate concentration would be 1000 times carbonate concentration meaning that $10^{-2.45} = 0.0035$ moles/L (216 mg/L) bicarbonate can be in solution with 40 mg/L calcium provided the pH is at 7.38. If there is enough hydrogen ion present to convert carbonate and lower pH then dissolution of limestone can take place as described.

The source of the hydrogen ions is dissolution of carbon dioxide and subsequent dissociation (per Equation (2.1)) of the carbonic acid so formed:



There is an equilibrium constant associated with Equation (3.4) and it relates the concentrations of carbonic acid in solution and of carbon dioxide in air. Because the latter is a gas, its concentration is expressed in terms of its partial pressure (the pressure it would exert if it were the only gas present).

$$p \frac{[H_2CO_3]}{P_{CO_2}} = pK_H = 1.47 \quad (3.5)$$

or, applying p in reverse

$$[H_2CO_3] = P_{CO_2} 10^{-1.47} \quad (3.6)$$

where the pressure, P_{CO_2} is in atmospheres. K_H is called the *Henry coefficient* and has units of moles per liter per atmosphere. The value given is for 20°C.

CO_2 is found in the atmosphere at a partial pressure of from 0.03 to 0.05 atmosphere but the partial pressure of the gas in soil can be orders of magnitude greater than that because of the presence of respiring soil bacteria. Even so only tiny amounts of it dissolves but most of that converts to bicarbonate at mid pH (at $pH \approx 7$ is 4.5). The protons released by the conversion are the protons which are available to convert carbonate to bicarbonate. The carbonic which has been converted must be replaced in order to satisfy the equilibrium condition of Equation (3.5) and so more CO_2 dissolves.

Figure 3.1 diagrams the process just described. It shows the four reactions we have already discussed (Equations (2.1), (2.6), (3.1), and (3.4)). There is a fifth reaction which reflects the fact that water is in fact an acid (and a base):



Its equilibrium constant dictates that

$$pH + p[(OH)^-] = pK_w = 14 \quad (3.8)$$

(at room temperature).

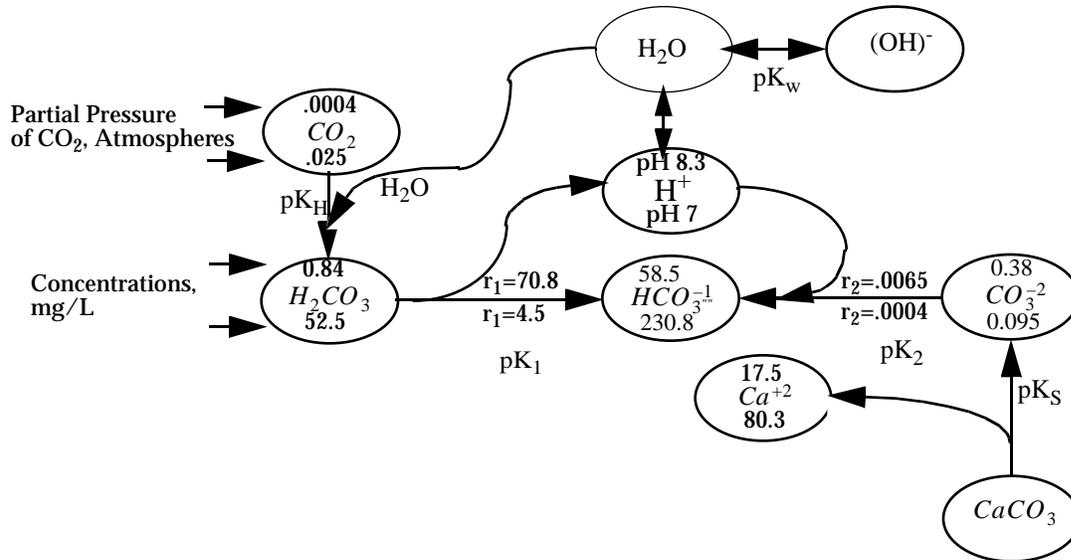


Figure 3.1 Carbo system in contact with limestone and carbon dioxide at two partial pressures of CO₂ showing flow of species as limestone dissolves. All five reactions discussed in the text take place simultaneously. Small numbers in ovals are concentrations in mg/L at the two pressures (indicated in CO₂ oval).

All five of these reactions proceed with carbon dioxide dissolving at one end and calcium carbonate at the other until a sufficient amount of each has dissolved that all 5 equilibrium conditions are met meaning that Equations (2.3), (2.7), (3.2), (3.5) and (3.8) are simultaneously satisfied. A sixth condition must also be satisfied and that is that the solution must be electrically neutral:

$$2[Ca^{+2}] + [H^+] - [HCO_3^-] - 2[CO_3^{2-}] - [OH^-] = 0 \quad (3.9)$$

Figure 3.1 (really the mathematical relationships it defines) shows that the parameter which defines the whole system is the partial pressure of carbon dioxide. The diagram indicates the concentrations of all quantities for two pressures namely 0.0004 atmospheres, which is typical of the partial pressure of carbon dioxide in the air and 0.025 atmospheres which is a level that can easily be achieved in soil with respiring bacteria. Thus surface waters (exposed to atmospheric CO_2) could be expected to have hardness (17.5 mg/L or 43.8 ppm as CaCO_3) and alkalinity (58.5 mg/L bicarbonate corresponding approximately to 48 ppm as CaCO_3) and they do. Conversely water drawn from a well and thus exposed to higher partial pressure of CO_2 can be expected to be harder and more alkaline and it is.

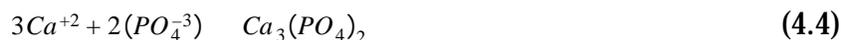
What we have described above all assumes that the water is in chemical equilibrium. Nothing has been said as to whether any particular water is in equilibrium, how it got to that state and if it is how long it took to get there. This is another subject altogether. The last mentioned well water is interesting in this regard. If while it was in the earth it was at equilibrium with CO_2 at a partial pressure 10 or 100 times what it would be at the surface it clearly isn't any more when brought to the surface. As it contains 10 or 100 times more CO_2 than it would in equilibrium with the air it is said to be *super saturated* with respect to CO_2 and will tend to lose CO_2 to the air until new equilibrium is reached. The situation is analogous to what happens when a beer is opened though as the amount of dissolved CO_2 is much less than that in a beer the rate of loss will not be as dramatic (bubbles will not form). Subterranean water contains not only more CO_2 but also much more calcium and bicarbonate than it would if it had dissolved them from open air levels of dissolved CO_2 . It is, thus, said to be oversaturated with respect to calcium carbonate as well and will tend to precipitate some if conditions are proper. Figure 3.1 with all the arrow heads reversed describes such a water and suggests that driving off carbon dioxide will result in removing bicarbonate from water and calcium with it.

Getting carbon dioxide out of the water can be hastened by heating the water and sparging it with air or steam. As CO_2 leaves the water, some bicarbonate converts to carbonic acid to replace the carbonic acid lost to CO_2 . This consumes hydrogen ions thus raising the pH. This causes bicarbonate to convert to carbonate which coalesces with calcium and precipitates. This causes more bicarbonate to convert to carbonate to replace it which releases hydrogen ions to combine with bicarbonate producing more carbonic which in turn converts CO_2 to and leaves the solution as gas. Thus boiling carbonaceous water which contains calcium (such waters are said to possess *temporary hardness*) is a time honored way of decarbonating it. This method is not very popular today because of the energy required.

The general strategy for decarbonation is to raise the pH in the presence of calcium. This is really how the boiling strategy works. Driving off CO_2 consumes hydrogen ions i.e. it removes acid from the solution. The resulting upward pH shift converts bicarbonate to carbonate which precipitates with calcium. Supplementing the calcium is a good way to make the method more effective. Other strategies for raising the pH are also effective. Various bases have been used to raise pH. Perhaps the most popular for brewers is calcium hydroxide (lime) or calcium oxide (quick lime) which converts to the hydroxide as soon as it contacts water. It simultaneously raises the pH and calcium concentration and can be used in cold water. This method is described in detail in Ref. 3. and elsewhere.

4.0 How Malt Phosphate Neutralizes Bicarbonate

Malt contains a fair amount of phosphate, in fact as much as 1% of its weight (as P_2O_5 - Ref. 4.). The term phosphate refers to compounds which involve the ions of phosphoric acid. This acid has 3 protons. Calcium phosphate is even less soluble in water than calcium carbonate and it is this fact which is behind the ability of malt phosphate to establish a favorable mash tun pH. To see how this works we need to add 4 more reactions to the system. These are



and each of has an equilibrium constant. They are, respectively, $pK_{p_1} = 2.12$, $pK_{p_2} = 7.20$, $pK_{p_3} = 12.44$ and $pK_{p_s} = 32.7$ (the P in the subscripts is used to indicate that we are describing the phosphate system). As with carbonic acid the relative quantities of the phosphate species are calculated from $r = 10^{(pH - pK)}$ except that this time there are 3 r 's rather than just two. One can also calculate the concentration of each species as a fraction of all the molecules containing phosphorous. This time $d = 1 + r_1 + r_1r_2 + r_1r_2r_3$ and there are 4 fractions including $f_4 = r_3f_3$ for the PO_4^{3-} ion.

The saturation condition for calcium phosphate is

$$3p[\text{Ca}^{+2}] + 2p[\text{PO}_4^{3-}] = pK_{p_s} = 32.7 \quad (4.5)$$

which indicates that it is quite insoluble. If we note that $[\text{PO}_4^{3-}] = f_4P_T$ where P_T is the total molar concentration of all species containing phosphorous we can use that in Equation (4.5) to see how much total phosphate can be in solution as a function of the amount of calcium in the liquor.

Doing this and taking the inverse of all the p 's gives

$$P_T = \frac{1}{f_4} \sqrt{\frac{10^{-32.7}}{[Ca^{+2}]^3}} \quad (4.6)$$

It is interesting to calculate some values at a pH within the range we would like our strike pH to fall. pH 5.5 is in this range. At that pH f_4 has a value of about 2.2×10^{-9} . If the calcium hardness of the liquor is 1 ppm as $CaCO_3$ then P_T has a value of about 650 millimoles/Liter. If the hardness is 5 ppm P_T is about 58 mMol/L, for hardness of 8 ppm it is about 30 mM/L and for 100 ppm it is 0.65. Earlier we said that malt contained about 1% phosphate (as P_2O_5) by weight. If the mash thickness is 1 quart per liter that means that the malt introduces 4.54 grams (as P_2O_5) per liter which is about 0.03 mole (a mole of P_2O_5 weighs 142 grams. Each mole of P_2O_5 reacts with water to produce 2 moles of phosphate so the amount of phosphate in a 1 pound per liter mash is about 0.06 or 60 mM/L. This much can be in solution at pH 5.5 as long as the calcium hardness of the water is less than 5 ppm. If the water is harder than this, then calcium phosphate will precipitate causing HPO_4^{-2} to convert to phosphate to replace the loss. In turn $H_2PO_4^{-1}$ converts to HPO_4^{-2} and even some phosphoric acid, H_3PO_4 converts to $H_2PO_4^{-1}$ to replace the $H_2PO_4^{-1}$ ions that converted to HPO_4^{-2} . Each of these conversions releases hydrogen ions and those hydrogen ions can react with carbonate and bicarbonate as discussed previously. If enough calcium is present it is possible to lower the pH to the point where most of the bicarbonate converts to carbonic and escapes the solution. The net process is depicted in Figure 4.1

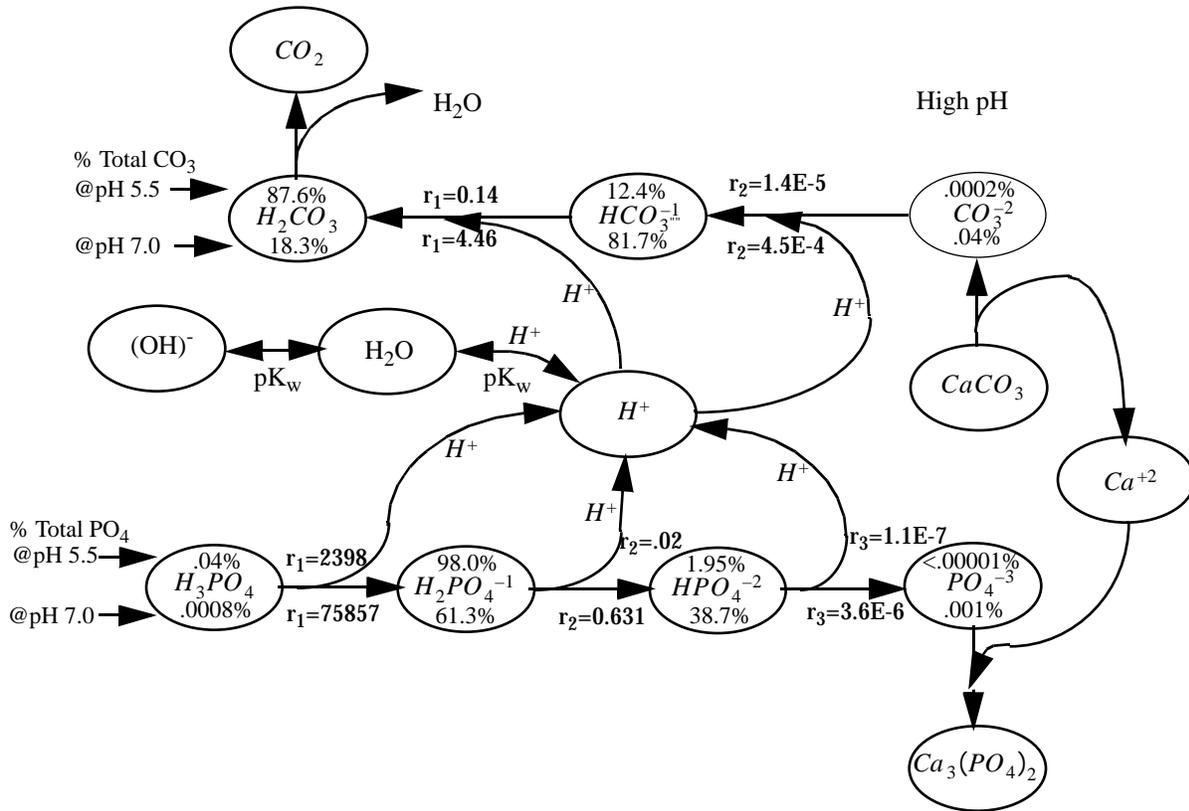


Figure 4.1 Conceptual diagram summarizing 9 chemical reactions involved in acidulation of mash by malt phosphate. Percentages in ovals are fraction of total carbonate or phosphate in indicated protonation state. Small numbers near arrows indicates molar ratio of less protonated to more protonated species. For example at pH 7 18.3% of xCO_3 is carbonic and bicarbonate moles are 4.46 times carbonic moles.

5.0 Kolbach's Residual Alkalinity

While it is possible to solve the 10 equilibrium equations implied by Figure 4.1 it is impractical and unnecessary. What is really important are the practical implications of the theory of the previous parts of the article. Paul Kohlback (Ref.5.) observed that in mashes consisting of pale malt, each 3.5 milliequivalents/L of calcium could “neutralize” 1 milliequivalent/L of water alkalinity. Seven milliequivalents of magnesium were required to do the same thing. Neutralize in this context means produce a mash pH that is the same as would be realized if the same malt were mashed with distilled (0 alkalinity) water. Alkalinity beyond that which could be neutralized by calcium and magnesium he termed “residual alkalinity”. As a simple formula

$$RA = Alk - \frac{\{Ca^{+2}\}}{3.5} - \frac{\{Mg^{+2}\}}{7} \quad (5.1)$$

The curly brackets mean that the concentrations can be in milliequivalents per liter or ppm as $CaCO_3$ (50 times the milliequivalents per liter value). The residual alkalinity will then be in the same units

If we give half the magnesium hardness plus the calcium hardness a name such as *effective calcium hardness* or just *effective hardness* and assign a symbol to it, say, $D = \{Ca^{+2}\} - \{Mg^{+2}\}/2$, then $RA = Alk - D/3.5$. We find it useful to rearrange this to $Alk = RA + D/3.5$ which plots on a graph of alkalinity vs. effective hardness as a straight line with slope 1/3.5. A series of such lines for various values of residual alkalinity is shown on Figure 5.1. Also noted on this plot, at coordinates corresponding to the approximate reported effective hardness and alkalinities of their water supplies, are the names of several brewing cities of renown (some cities appear more than once because there are multiple reports as to the compositions of their waters). Figure 5.1 is thus a convenient tool for comparing one's own water supply to the water of these cities.

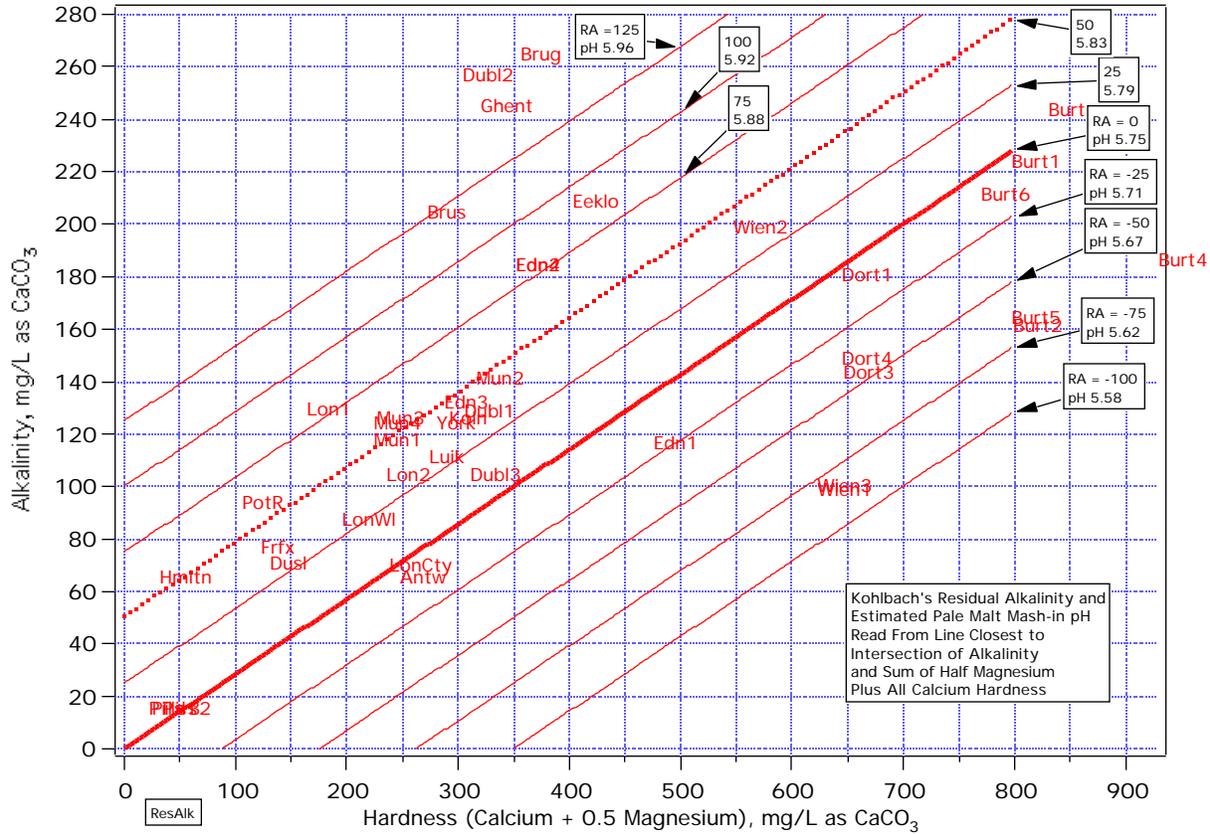


Figure 5.1 Residual alkalinity and mash pH prediction chart

A second finding of Kohlbach was that each unit of residual alkalinity (as CaCO_3) shifts the pH relative to a distilled water mash by .00168. Thus a city with water yielding a RA of +100, such as Eeklo (Netherlands) would be expected to have mash pH 0.17 units higher than a distilled water mash. Assuming this to be 5.75 would imply 5.92 with the +100 RA water which is clearly too high. By contrast a water with RA of -75 would yield a mash tun pH of about 5.62 which is acceptable. Vienna, with its very hard water and Burton with its extremely hard water would thus both give an acceptable strike pH without the use of external acid.

Readers can, of course, plot their own water's coordinates on Figure 5.1 and by so doing get a rough idea of what pH a base malt only grist might produce at strike. If this is unacceptable, the approximate amount of acid required to set the pH to a desired value is easily estimated. Assume, for example, water like Munich's with an RA of 50 leading to a mash pH with base malt of about 5.83. The RA can be shifted by adding 1 mEq/L of acid per liter of liquor for each 50 units of shift desired. Thus the Munich water can be brought to 0 RA by the addition of 1 mEq/L of acid. This would result in a mash pH of about 5.75. If it is desired to have a lower pH, say pH 5.67, the RA would need to be brought to -50 and another mEq/L would be required. These 2 mEq/L of acid could be supplied by adding 2 mL of 1 N acid to the mash. It is also quite possible to obtain this acid from dark malt. The amount of dark malt required can be calculated if the malt's titratable acidity is known.

We would be remiss if we did not point out that residual alkalinity and the implied pH shift represent approximations of what will actually happen in the mash tun. Given these calculations should be used only as a starting point in the design process. Actual dark malt or mineral acid additions should be determined by experimentation with test mashes.

6.0 Sidebar: Review of Water Chemistry Terms

6.1 Mole

A mole is 6.02×10^{23} atoms, ions, molecules, oranges or bowling balls. In chemistry it is a fundamental measure of the amount of a substance. To convert amounts expressed as moles to grams multiply by the gram molecular weight. For carbon dioxide this is 44, for carbonic acid, 62, for bicarbonate ion, 61, for carbonate ion 60, for calcium ion 40. To convert from grams to moles divide by the gram molecular weight. A millimole is 1/1000 of a mole.

6.2 Equivalent

An equivalent is the amount of an ion that carries one mole of charge. The equivalent weight of an ion is its molecular weight divided by its charge (the sign is ignored). For example, carbonate ion with a molecular weight of 60 and a charge of -2 has an equivalent weight of 30.

6.3 Alkalinity:

Alkalinity is a measure of the water's ability to resist acidification. It is measured (Ref. 6.) by adding, incrementally, a standard solution containing 0.1 equivalent of acid per liter to a 100 mL sample of the water. The number of mL of the standard required to bring the pH to 4.3 is the alkalinity of the water expressed in milliequivalents per liter (mEq/L). Multiplying this number by 50 gives the alkalinity in ppm as CaCO₃.

In typical brewing water alkalinity is attributable almost entirely to the presence of the bicarbonate ion. If bicarbonate ion concentration is known in mg/L it can be divided by 61, the gram molecular weight of bicarbonate to give the number of milliequivalents per liter of alkalinity. As above, multiplying by 50 gives the more usual units of ppm as CaCO₃.

6.4 Hardness

Hardness is a measure of the water's calcium and magnesium contents. To measure the total hardness an indicator which is of different color dependent on the presence or absence of free calcium or magnesium (e.g. Eriochrome Black T is red in the presence of calcium or magnesium ions and blue when they are absent) is added to 100 mL of sample and then a solution of a chelating agent (e.g. EDTA) of known strength, is added incrementally until the indicator changes from red to blue. If the EDTA is, for example, capable of chelating (binding) 0.1 equivalent of calcium or magnesium per liter then the hardness (in mEq/L) is the number of mL added to 100 mL of sample to bring about the color change. Multiplying by 50 gives the hardness as CaCO₃. To obtain separate values for the calcium and magnesium hardnesses, one or the other is removed from the sample (e.g. by precipitating it) and the hardness measured again. The new measurement gives the hardness of the component not removed and subtracting this value from the total hardness gives the hardness of the other component (Ref. 7.).

If calcium ion concentration is known in mg/L dividing by 20, the gram equivalent weight of calcium, gives the mEq/L calcium hardness. For magnesium, the gram equivalent weight is 12.5. Conversely multiplying the hardnesses as mEq/L by the equivalent weights give the hardnesses on mg/L.

6.5 Residual Alkalinity

The sum of the water's calcium hardness and half its magnesium hardness subtracted from the water's alkalinity. See text.

6.6 Ion

An ion is a charged particle formed when a neutral atom or molecule either yields up or acquires negatively charged electrons or positively charged protons.

6.6.1 Cation

A cation is a positively charged ion usually formed when a metal atom gives up one or more electrons. In water the commonly found cations are those of hydrogen, calcium, magnesium, sodium, and potassium. In some supplies iron, copper, manganese, strontium and other metal ions are found as well. In some samples the ammonium ion, formed when a molecule of dissolved ammonia takes up a proton, will be found.

6.6.2 Anion

An anion is a negatively charged ion. These can be formed when an atom such as chlorine acquires an electron or when an acid, such as sulfuric acid, sheds its protons. Ions commonly found in water are chloride, sulfate, nitrate, and bicarbonate. Less commonly found examples are nitrite, and carbonate.

6.7 pH

pH is a measure of the balance between alkalinity and acidity in a sample. It is approximately equal to the negative logarithm of the hydrogen ion concentration. Solutions with an excess of acid have a $\text{pH} < 7$. Those with an excess of alkali, $\text{pH} > 7$

6.8 Titrateable Acidity

Titrateable Acidity is the amount of acid a malt releases in lowering pH to the target mash pH. It is determined by adding a ground sample of the malt to distilled water and then titrating with (adding measured quantities of) a base of known strength until the pH reaches the target. The number of milliequivalents of base used divided by the weight of the malt is the titrateable acidity of the malt in units of milliequivalents per pound or milliequivalents per kilogram.

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