# Accurately preparing large quantities of concentrated hydroponic nutrients

When preparing concentrated solutions for hydroponics it is important to have a reproducible process that always generates the exact same results. If this is not done, you'll obtain different nutrient concentrations between different batches and the concentrated nutrient additions to create the final nutrient solutions will yield inconsistent results. To address the potential variability of the concentrated solution manufacturing process we need to understand the different sources of error present and come up with ways to modify the process to generate more reproducible results. In this blog post I will talk about the largest source of error when preparing larger batches of concentrated nutrient solutions and how this error can be greatly reduced in order to obtain both more precise and accurate results.



Picture of a type A 250mL volumetric flask.

The process of preparing hydroponic concentrated solutions involves two steps. First, you dissolve raw fertilizer salts into some volume of distilled or RO water and then you take this volume of solution to a desired final volume of solution using the same source of water. In a small scale setup this process is very simple to carry out, since we can just weight and dissolve all our salts in some fraction of the desired final volume and then use a precise instrument to measure total volume - most typically a volumetric flask - to take our solution to the final desired volume. For example if we desire to prepare 250 mL of concentrated nutrient solution and we use a well calibrated scale with +/-0.001g of precision and an A grade volumetric flask with a precision of  $\pm$  0.3mL, the error we expect to get from a 500mg salt will be +/- 4.77 ppm with a 99% confidence. Since the concentration of this salt in the concentrated solution is 2000 ppm, we get a final result of 2000 +/- 4.77 ppm. If both instruments are calibrated this is a very precise and accurate result.

When we move to larger amounts of solution we usually get better on the side of mass. This is because we can still get scales that weight with +/-0.1g precision even at weights exceeding 50kg, so our error as a fraction of the total measurement remain in the 0.01% to 1% region pretty easily. However things get way worse in terms of volume. If you are preparing 100 gallons of nutrient solution — around 378 liters — you will be able to weight the salts precisely and accurately but when it comes to measuring final volumes of solution, you are not going to be very lucky. The volume marks in tanks are widely inaccurate and are not even standardized to any level of significant precision or accuracy plus accurately measuring whether water is at a given level in a tank is a very error prone process because of how wide the tank area is.

Although we don't usually have a way to adequately measure final volume, we do have a way to measure volume going into a

tank in the form of flow meters, which can give us significant accuracy and precision. However, to be able to properly use the flow meter — know how much volume we need to actually get to the final volume we want — we must obtain information from a precise and accurate low scale process. To do this you can carry out the following steps:

- Get a precise and accurate scale (calibrated and at least +/- 0.001g in precision)
- Get a scale that can weight up to 500g that can measure with at least +/- 0.1g precision (if the one above does not).
- Get a 250 mL type A volumetric flask (should be around +/- 0.3 mL in precision).
- Get a 250mL beaker
- Get a plastic lab washing bottle and fill it with distilled water
- Calculate the salts you would need to dissolve to arrive at your desired concentrations at a 250mL final volume of concentrated solution
- Weight those salts and put them in a beaker, take note of all the exact weights added.
- Weight the dry, empty volumetric flask
- Add approximately half the volume of distilled water to the beaker and dissolve the salts
- Transfer to the volumetric flask, use the washing flask to fill the volumetric flask up to the calibration line (bottom of water meniscus is touching the line when viewed at eye level).
- Weight the flask with the solution
- Calculate the weight of water (weight of flask with solution – weight of flask – sum of weight of salts)

If the procedure above was carried out between 10-25C (50-77F) we can approximate the density of water to 1.0g/mL with little error (around 0.003g/mL). This means that we know the volume of water that was required to get to the desired final volume

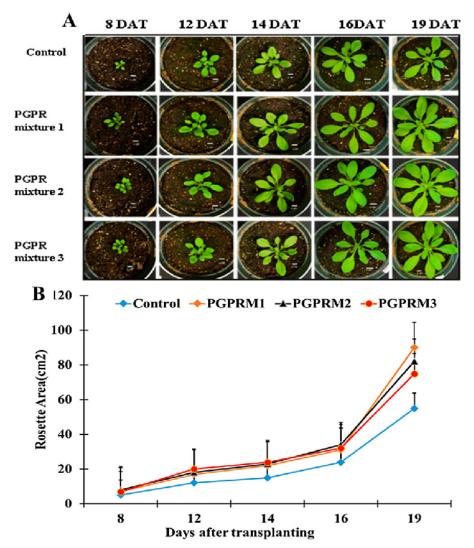
and we can then transfer this volume to our preparation procedure when we use a large tank. If the volume of water required for the preparation of the 250mL solution was just 230mL, then we can assume that the volume required to prepare 100 gallons will be 92 gallons, as the salts, when proportionately scaled, will take up the same volume and will require the same amount of water proportionately to reach the final desired volume.

When this type of procedure is done and an accurate and precise flowmeter is used, we can usually achieve concentration values at large scales that will be in the 0.1-1.0% error range, which is way better than anything that can be achieved by just using lines in tanks or procedures that use flow meters but ignore what the actual amount of water added needs to be in order to reach the desired concentration (many people achieve the salts take up no volume, which is a mistake). Having low errors in concentrated solutions means there will be less variability in final nutrient solution composition and therefore more reproducibility in crops.

# Plant Growth Promoting Rhizobacteria (PGPR) in hydroponics

Plants did not evolve in an isolated environment but with a wide variety of different microbes. Through their evolution, plants prospered more in the presence of certain microbes and therefore evolved traits to attract and nurture them. In turn these microbes were also selected to create even deeper

mutualistic relationships with plants. Specifically, the bacteria from this group that facilitate and improve plant growth are known as Plant Growth Promoting Rhizobacteria (PGPR) and have been an extensive subject of plant research during the past 40 years. In this article I am going to talk about their use in hydroponic culture and the evidence we have about their growth promoting effects in the absence of soil.



Effect of PGPR of the genus Bacillus in soil, taken from <a href="mailto:this">this</a>
<a href="mailto:paper">paper</a>

The positive effects of PGPR in general are well established. These two  $(1,\ 2)$  literature reviews address the subject in depth and cite a lot of the research that has been done around PGPR for crops in general, although none of these two reviews address their use in hydroponics specifically. What we know from all these literature is that the positive effects of PGPR are mostly attributed to three different phenomena. The first

is an increase in nutrient availability for the plant, mainly through making some nutrients that are inaccessible to the plant accessible (mostly N and P), the second is through the release of phytohormones — chemical substances that stimulate plant responses — that prompt plants to develop more tissue in several different ways, and the third is that these bacterial colonies provide defenses against pathogens that could be attacking the plant if they were not present. Many different species that show these effects have been identified — some even specific to single plant species — but from those species those from the genus Bacillus, Agrobacterium and Pseudomonas have been the most widely studied and shown to be effective.

We also know from the research that the application of PGPR is not trivial and exactly how plants are inoculated with them plays an important role in the improvements they might show. Inoculation can be done in seeds, cuttings, transplants or through the entire growing/flowering periods. You can use both root and/or foliar applications, different concentrations of bacteria and different additives can also be given to try to make the inoculation steps more successful. These bacteria can also use oxygen in solutions, so using too much can also starve roots of important oxygen and cause strong negative effects before any positive effects can be seen, using too little means the bacteria die without being able to form a stable colony. The table below gives you an idea about how complex the entire application universe can be and the sort of effects that have been observed in field/greenhouse trials in soil for a wide variety of plants. The reviews cited above contain a lot of additional references, make sure to read them if you're interested in a wider view of the available literature on the subject.

Table 2
Effects of plant growth-promoting rhizobacteria (PGPR) application on fruit crops.

Crop	PGPR (species/strain)	Application mode	Experimental conditions	Effects	References
Apple	Bacillus sp. strain M3 <sup>a</sup> and OSU-142 <sup>a</sup> , Microbacterium sp. strain FSO1 <sup>e</sup> , Pseudomonas sp. strain BA-8 <sup>d</sup> (alone or in combinations)	Root-dipping (10 <sup>9</sup> CFU mL <sup>-1</sup> )	Field	Increased cumulative yield, fruit weight, shoot length, and shoot diameter in apple cv. Granny Smith and Stark Spur Golden	Karlidag et al. (2007); Aslantas et al. (2007)
	Bacillus sp.a	Foliar application of	Field	Enhanced growth of apple leaves and improved fruit	Ryu et al.
Apricot	Bacillus sp. strain OSU-142a	spores (10 <sup>7</sup> spores g <sup>-1</sup> ) Foliar application (10 <sup>9</sup> CFU mL <sup>-1</sup> )	Field	quality parameters (sweetness and moisture content) Increased yield, shoot development and reduced shot-hole disease severity and incidence	(2011) Esitken et al. (2002, 2003)
anana	Pseudomonas fluorescens strain CHAO <sup>d</sup>	Soil application of cells (2.5-3 10 <sup>10</sup> CFU) with or without chitin (treatment repeated three times)	Field	Increased growth, leaf nutrient contents and yield of banana plants under perennial cropping systems	Kavino et al. (2010)
Cherry	Pseudomonas sp. strain BA-8 <sup>d</sup> and Bacillus sp. strain OSU-142 <sup>a</sup> (alone or combinations)	Foliar application (spray; 10 <sup>9</sup> CFU mL <sup>-1</sup> )	Field	Stimulated plant growth, increased yield per trunk, fruit weight and shoot length and resulted in significant yield increase	Esitken et al. (2006)
Grape	Pseudomonas putida strain BA-8 <sup>d</sup> and Bacillus simplex strain T7 <sup>a</sup> (alone or combinations)	Grafted plant-dipping (10 <sup>9</sup> CFU mL <sup>-1</sup> ) for 60 min	Experimental glasshouse	Increased graft callusing, scion shoot growth, cane hardening, and nursery survival rate, as well as fruitfulness of the grapes in following year	Sabir (2013)
Hazelnut	N <sub>2</sub> -fixing and P-solubilizing bacteria	Seed-dipping (10 <sup>9</sup> CFU mL <sup>-1</sup> ), on one-year old seedlings	Pots, greenhouse conditions	Increased seedling and total branch length, branch number, trunk diameter, and nutrient uptake	Erturk et al. (2011)
Kiwifruit	Bacillus sp. <sup>a</sup> , Paenibacillus polymyxa <sup>a</sup> and Comamonas acidovorans <sup>c</sup>	Seed-dipping (10 <sup>9</sup> CFU mL <sup>-1</sup> ) for 30 min	Greenhouse conditions	Stimulation of rooting and root growth	Erturk et al. (2010)
Strawberry	Bacillus subtilis strain GBO3ª and Bacillus amyloliquefaciens strain IN937aª	Seed-dipping with a formulation that contains both strains in a 2.5% chitin carrier	Field	Addition of PGPR to plug transplants resulted in healthier roots, earlier and higher total yields	Kokalis-Burelle (2003)
	Bacillus sp. FS-3 <sup>a</sup>	Root drench (3.5 × 10 <sup>7</sup> cell g <sup>-1</sup> ), repeated five times within 7-D intervals	Pots, greenhouse conditions	Increased fruit and leaf nutrient concentrations (N, P, K, Ca, and Fe)	Güneş et al. (2009)
	Azospirillum brasilense strain REC3 <sup>b</sup> , RLC1 <sup>b</sup> , PEC5 <sup>b</sup>	Root-dipping (10 <sup>6</sup> CFU mL <sup>-1</sup> ) for 30 min	Pots. greenhouse conditions	Increased root length, root area, and dry weight of root and shoot	Pedraza et al. (2010)
	Pseudomonas sp. strain BA-8 <sup>d</sup> and Bacillus sp. strain OSU-142 <sup>a</sup> and M3 <sup>a</sup> (alone or combinations)	Root-dipping (10 <sup>9</sup> CFU mL <sup>-1</sup> ) for 30 min or foliar application	Field	Increased fruit yield, plant growth, phosphorus and zinc content of leaves	Esitken et al. (2010)
	Bacillus spharicus GC subgroup B strain EY30 <sup>a</sup> , Staphylococcus kloosii strain EY37 <sup>a</sup> and Kocuria erythromyxa strain EY43 <sup>e</sup>	Root-dipping (10 <sup>8</sup> CFU mL <sup>-1</sup> ) for 30 min	Pots, greenhouse conditions	Increased plant growth, fruit yield, chlorophyll content, relative water content of leaves, mineral uptake (N content of leaves and P content of roots), and reduced membrane injury under saline conditions (35 mM NaCl)	Karlidag et al. (2010)
	P. fluorescens strain Pf4 <sup>d</sup> , Pseudomonas sp. strain 5Vm1K <sup>d</sup>	Root drench with the two PRGB (5 10 <sup>9</sup> CFU) and/or with arbuscular mycorrhizal fungi	Pots, greenhouse conditions	Increased anthocyanin concentration in fruits of plants grown under conditions of reduced fertilization	Lingua et al. (2013)
Walnut	Alcaligenes sp. strain 637Ca <sup>c</sup> Pseudomonas chlororaphis <sup>d</sup> Arthrobacter pascens5	Root-dipping (10 <sup>8</sup> CFU mL <sup>-1</sup> ) for 30 min Seed-dipping (10 <sup>9</sup> CFU mL <sup>-1</sup> ), on one-year old seedlings	Pots, greenhouse conditions Greenhouse conditions	Increased fruit yield, number and weight under high calcareous soil conditions Increased plant height, shoot and root dry weight, phosphorus and nitrogen uptake	Ipek et al. (2014) Yu et al. (2012)

Table showing the effects of different PGPR applications using different techniques across different plants. Taken from <a href="this:review">this</a> <a href="teview">review</a>.

As you can see the effects under these conditions have been very positive, with sometimes highly significant increases in root/shoot weights and fruit/flower yields. However soil itself is not a perfect media and plants grown in soil are also not subjected to ideal nutrition. Since one of the main benefits of PGPR is to increase nutrient availability, some of these benefits might be partially or even completely negated when moving onto hydroponic culture, where we seek to provide

plants with an ideal environment. Research of PGPR in hydroponics is not very common though, as hydroponic growing has traditionally made a big deal about sterility, as growers mostly want to prevent pathogens from getting into their crops.

Ref	Plant	PGPR	Yield	Link
1	Tomato	Pseudomonas fluorescens, Pseudomonas putida	10%+	https://www.actahort.org/books/952/952_98.htm
2	Tomato	Pseudomonas fluorescens	13%+	https://www.sciencedirect.com/science/article/abs/pii/003807179390038D
3	Tomato	Pseudomonas putida, Serratia marcescens, Pseudomonas fluorescens, Bacillus spp	18-37%+	https://www.actahort.org/books/807/807_68.htm
4	Cucumber	Pseudomonas putida, Serratia marcescens, Bacillus spp., Pseudomonas fluorescens	78-121%	https://www.sciencedirect.com/science/article/abs/pii/S0304423813000198
5	Tomato	Bacillus amyloliquefaciens	8%	https://dialnet.unirioja.es/servlet/articulo?codigo=2740834

References of some trials using PGPR carried out in hydroponic conditions

Thankfully there have been some people who have led the way into the world of PGPR in hydroponic research so we have started to see some positive evidence of their use, even under hydroponic growing conditions. The above table shows you 5 references for papers that have studied PGPR in hydroponics — mainly in tomato plants — where it has been pretty well established that applications of bacteria of the genus Pseudomonas can increase yields in the order of at least 10%+. Some studies, like 3 and 4, show that significantly more gains are possible for different combinations of bacteria or application methods. I couldn't find a lot of additional studies in this direction, but the above studies start to show that the use of these bacteria in hydroponics can be positive.

A lot of questions still remain though. If these bacteria are benefiting plants because of the introduction of plant growth regulators (PGR) in solution, then we might ask if the direct exogenous applications of these PGRs is not a better way to obtain and control the benefits without the need to maintain a live population of bacteria in a mutualistic relationship with plant roots. Research has indeed shown that the exogenous application of many PGRs can enhance the yields of different plants. Do we apply PGRs or do we keep a culture of bacteria in our media? Can we do both and obtain even better results? Sadly right now there are no answers to the above questions and a lot of additional research is needed before we even get close.

For now the research on PGPR is telling us that these bacteria work amazingly well in soil and can also provide substantial benefits for some plants in hydroponic culture under certain conditions. We know that the bacteria from the genus Pseudomonas and Bacillus are the most interesting candidates to study in hydroponics and we know some of the inoculation techniques that have worked. If you want to experiment with them in your hydroponic crops, make sure you take the above information into account. The right choice of bacteria, concentration, inoculation method and additives can make a big difference in the results you get.

#### Why do NPK labels express P and K as oxides?

If you have had any contact with the fertilizer world you have probably noticed that fertilizer labels contain N-P-K values on their front and back labels, denoting the chemical composition of nitrogen, phosphorus and potassium available within the product. However you will soon learn that while N is elemental composition — the actual percent of the

fertilizer by weight that is nitrogen — P and K are expressed in more confusing terms, mainly the oxides  $K_2O$  and  $P_2O_5$ . Why do we keep expressing these elements as oxides? Is there any actual reason why expressing them as oxides would be better? What's the point? In today's post we'll talk about fertilizer and fertilizer analysis, we'll talk about why P, K and other elements are expressed as oxides and why this continues to be the case.



Nitrogen, phoshprous and potassium are the elements represented in the N-P-K, although P and K are expressed as oxides and not pure elemental forms

I have heard people talk about the expression of K as  $K_2O$  and P as  $P_2O_5$  as a consequence of K and P not being actually present in their elemental forms in the fertilizers but as other substances. The argument being that it is preferred to express these elements as their available forms, instead of their elemental forms. However this argument has many problems. The first is that  $K_2O$  and  $P_2O_5$  are also not present within the fertilizer, as these two are also very reactive forms of these elements. Potassium in particular is always present as  $K^+$  ions, reason why it would make more sense to express it as elemental potassium and P is actually present most commonly as either

 $H_2PO_4^{-2}$  or  $HPO_4^-$ , all of these pretty far away from the phosphorus pentoxide form that the label describes it as  $(P_2O_5)$  is *not* phosphate). Nitrogen is also not present as elemental N, but it is present most frequently as either  $NO_3^-$  or  $NH_4^+$  ions (although urea and amminoacids are also common forms of N in non-hydroponic fertilizers).

Why is N expressed as elemental N and K and P are not? The reason has to do with the way that these elements were quantified in the past when doing chemical analysis. Before we had access to modern techniques — such as inductively coupled plasma mass spectrometry — the elements were quantified using more complicated analysis procedures. The nitrogen was usually quantified using methods such as Kjeldahl nitrogen analysis because it would become volatile when the sample was burned, while the other elements were quantified from a calcined sample, meaning the sample was exposed to high temperatures to eliminate all water and carbon within it before the analysis. This ash would contain all non-volatile elements and when determining K and P from these ashes you could sometimes actually quantify  $K_2O$  and  $P_2O_5$ . From an analytical chemistry perspective, it made sense to express all non-volatile elements as oxides, because the concentration of these oxides was what you were actually measuring in the lab after you calcined the sample. This practice was very common in inorganic chemistry in general, because analysis of many nonvolatile elements tended to follow a similar path. The above is certainly an over-simplification, you can read more about analytical methods used in the early days of fertilizers here, if you do so pay special attention to the references in that paper.

In the past knowing the composition of fertilizers expressed in this way made sense, as labs could basically eliminate an additional conversion step when reporting and comparing results. Note that in those days — 1930-1950 — there were no

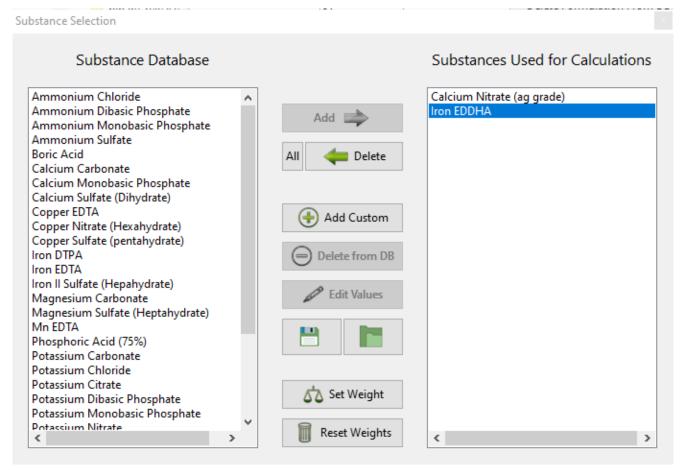
pocket calculators and everything needed to be calculated entirely by hand, so saving calculation steps was considered less trivial than it is right now as someone would actually need to make all those conversions using pen and paper. If you have to analyse 30 fertilizer samples in your lab then you would rather report a number closer to the one you directly measured instead of having to do 30 additional calculations by hand to get to another number. Since all labs were measuring these elements in similar ways, everyone agreed that it made sense for fertilizer labels to be  $N-K_2O-P_2O_5$ .

We no longer do things this way, as the methods and tools available to the analytical chemist have changed through time, but we keep this trend of reporting things in this manner in order to have coherence with past NPK labels. We have measured NPK in this manner for almost a century — the era of modern fertilizers starts in the early 1930s — so it would be a nightmare to change since it would become difficult to know when looking back which values were expressed as  $K_2O$  and  $P_2O_5$  and which ones as actual elemental P and K if the change was made.

So expressing K and P as  $K_2O$  and  $P_2O_5$  makes little sense in the modern world. We do it because we inherited this from the birth of the fertilizer era and we do it because making the conversion in these times is trivial and maintains coherence with all our previous reports of fertilizer compositions. However it is important to realize that  $K_2O$  and  $P_2O_5$  are not the actual forms that these elements have in fertilizers and that we simply express them this way through mathematical operations. Just image you're saying: "If the K present in this fertilizer was actually all  $K_2O$ , then it would be x% of the mass of the fertilizer".

#### HydroBuddy has now been updated to v1.70: New features and modifications

My free and open source hydroponic nutrient calculator has been available since 2010, going through many iterations and changes through the years. The latest version as of May-24-2020 is now 1.70, which you can download <a href="here">here</a>. This new release implements some important updates and modifications. In this post I will write about these, the reason why they have been made and the features that I am implementing for the next version of the software.

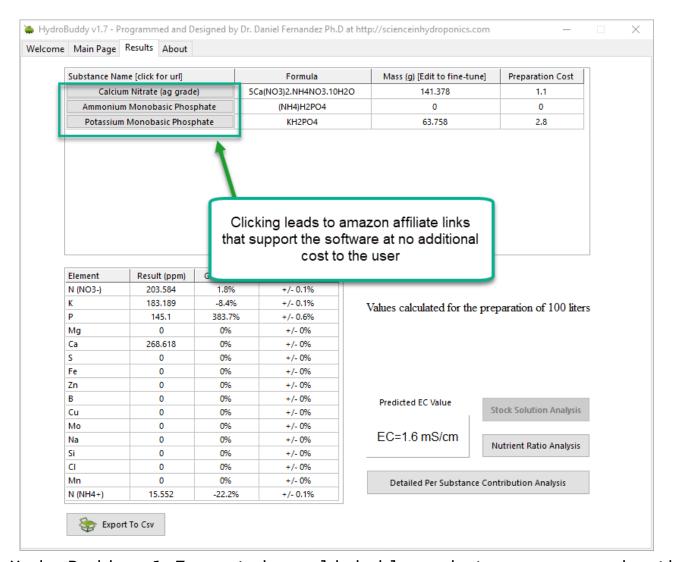


New substance selection screen in HydroBuddy v1.70

Most changes in this version have been done in the "Substance Selection" section of the program, which is accessible through the button of the same name in the "Main Page" tab. This is the "heart" of the program as this is where users decide what raw inputs they want to use and where they can manage the library of inputs that are actually available for calculations. In previous versions a very wide library of inputs was available by default, including many inputs that were rarely of any practical use in hydroponics and were there for illustrative purposes. A good example of this is a salt like "Calcium Nitrate (Tetrahydrate)" which is very rarely used by hydroponic growers as commercial "Calcium Nitrate" is actually a calcium ammonium nitrate salt that is very different in chemistry and composition to pure calcium nitrate tetrahydrate.

To solve the problem mentioned above I have completely rebuilt the substance database to include only commercially available raw fertilizers that make sense and are actually used in common situations in hydroponics. This included adding a lot of different metal chelates and salts that were previously ignored but are now part of the HydroBuddy default database.

Another issue I wanted to address was the confusion some users have about where to buy these chemicals and potentially get some revenue to support the development of the software at no additional cost to the user. For this reason I have added manually selected links to all the raw fertilizers that are included with the DB so that users who want to buy small quantities of those can also support the software when they do so.



HydroBuddy v1.7 contains clickable substance names in the result tab that take you to amazon affiliate links that sell the products mentioned at no additional cost to the user.

The "Substances Used" tab has also been enhanced with a new "Save/Load" functionality that enables users to save or load lists of substances used to avoid the hassle of having to go through and select substances whenever they want to prepare a certain solution. This has also been very annoying for me in the past as having to go through different sets of inputs used for different purposes can be a very time consuming exercise. With this new feature all I have to do is save one list for each one of my needs and a single click of the "Load" button can easily change a list of 5+ inputs without the need for any tedious and — mistake prone — manual changing. Another small manual enhancement has been the addition of a small "All" button next to the "Delete" button, which allows you to delete all the substances present in the "Substances Used for

Calculations" list.

Another change in this version was a decision to go with a 32 bit compiler in Windows in order to ensure that the variables for this operating system are all 32 bit. This will enable users who are using both 32 and 64 bit operating systems to use the software without problems. This was an issue in the past as many uses still use old 32 bit systems and they were having problems having to manually compile Hydrobuddy in some of their old machines. Sadly I still do not own a Mac, so HydroBuddy has yet to be available as a download for MacOSX and the software will need to be individually compiled by all of those who wish to use it in their MacOSX setups.

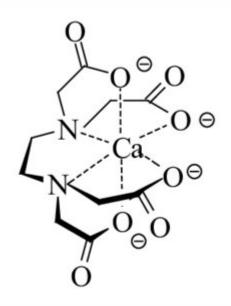
One of the features that is lacking most now is an ability to import databases from previous versions, as each time the software is updated users haven't been able to take advantage from previous custom databases built using the software due to problems with compatibility across releases (new DB fields being added, edited, etc). For the next version of the software I am working on a DB importing feature that should eliminate this issue so that users can benefit from the latest HydroBuddy releases without having to tediously add all their old substances to the new release.

With all the above said, I hope you enjoy this new version of the software. If you have any suggestions or comments about the above please feel free to leave your comments in this post!

#### Calcium EDTA and its problems

#### in hydroponics

Calcium is mainly used in hydroponics as calcium nitrate, given that this is a very soluble and abundant form of calcium. However this is not the only way calcium can be fed to plants and a myriad of other calcium sources exist. Among this we find calcium sulfate, calcium chloride, calcium hydrogen phosphate, calcium citrate, calcium gluconate and calcium EDTA. This last form, a chelate of calcium with EDTA, is one of the most cheaply available forms of chelated calcium but carries with it some substantial problems in hydroponic culture. In this article we are going to talk about Ca EDTA, its advantages and challenges when used as a supplement for calcium in hydroponics.



Model representation of the CaEDTA $^{-2}$  anion in the Ca EDTA salt. When talking about Ca EDTA we should first understand that this is not simply a calcium ion with an EDTA molecule wrapped around it. In reality, the product we purchase as Ca EDTA, that contains 9.7% Ca by weight, is actually represented chemically as  $C_{10}H_{12}O_8CaN_2Na_2\cdot 2H_2O$ . The Ca EDTA product is actually four parts, a few waters of crystallization, the  $Ca^{+2}$  cation, the chelating agent anion that wraps around it  $(EDTA^{-4})$ 

and two sodium cations, Na<sup>+</sup>, that are used to counter the two excess negative charges coming from the Ca EDTA (which we should more accurately call (CaEDTA)<sup>-2</sup>). When adding Ca EDTA we are actually adding four things, a little water, Ca, EDTA and Na. Most importantly Ca EDTA is in reality 12.15% sodium, meaning you're adding more Na than you're adding Ca when you use it.

Because of the above, thinking about Ca EDTA as any significant portion of a plants Ca nutrition is going to be a problem. Adding 100 ppm of Ca through this chemical would imply adding more than 100 ppm of Na. This addition of sodium can start to be heavily detrimental to plants as higher and higher values are reached (read my article on sodium in hydroponics to learn more). Although there is not much in the way of scientific literature using Ca EDTA, we do find some reports talking about heavy toxic effects at concentrations near 2.5 mM (940.7 ppm), which would contribute around 90 ppm of Ca to a solution.

Another important aspect to consider is the EDTA molecule itself. The EDTA chelate is not passive by any means and is not covalently attached to the Ca, so can easily move away. Since it binds pretty weakly with Ca, it will want to exchange Ca with anything else that seems more attractive to it. This poses an important problem when applying it in solution, as the EDTA in Ca EDTA might dissociate from Ca and attach to another ion that it finds more attractive, it prefers heavy metals so this can actually cause extraction of things like lead from the media. This might be an important consideration when used in cases where the media might contain significant amounts of heavy metals.

Yet another interesting issue — that I haven't seen mentioned anywhere else and only know experimentally — is that the actual CaEDTA<sup>-2</sup> anion can form insoluble salts with Ca itself. This means that you can actually precipitate Ca(CaEDTA) in

solutions that are highly concentrated in both ions. This is an important reason why concentrated solutions of Ca EDTA and Ca nitrate are very hard to prepare right, because as soon as you pass the solubility limit of Ca(CaEDTA) you will start to see it crystallize out of solution. Many people wonder why something is precipitating out of a solution made of two very soluble Ca salts, the reason is that Ca EDTA is not a neutral entity but can actually form a salt with free Ca. The Ca EDTA definitely requires its own concentrated solution most of the time.

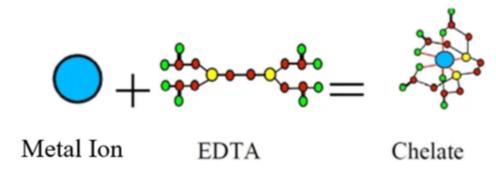
So why would anyone use CaEDTA given the above set of very important problems? There are a some advantages to it that make it a good salt for some applications, particularly foliar sprays. The first is that it is not going to precipitate easily out of solutions because of anions, so it can remain at a high concentration with anions that would normally precipitate as Ca salts in the presence of free Ca. This can be interesting in the case of some anions, like salicylates, that are often used as plant growth promoters (you can see this specific use in this paper). It is also one of the only forms of Ca that is taken in by the plant as an anion, so it is Ca that can get into the plant without having to compete with other cations in their transport channels. There are therefore some cases where Ca can be used very successfully in foliar applications (1).

Although there might be some niche applications for CaEDTA, particularly allowing some experiments that would be impossible with regular Ca salts, there are also some very important issues with its use in hydroponic culture. If you're contemplating using it, I would suggest you carefully consider its chemistry in solution and interactions with other substances that will be with it, particularly in stock solutions. You should also consider the amount of sodium being added and preferably avoid using it in feeding solution applications unless you have carefully considered all of the

above and its advantages are more important for your particular use case.

### How to prepare a low cost chelated micronutrient solution

Micronutrients constitute only a small portion of a plant's nutritional requirements but are still vital to growth and development. They are mainly comprised of heavy metals (Fe, Zn, Mn, Cu, Mo) as well as a single non-metal, boron (B). Since they are used in such small concentrations — normally in the 5 to 0.01 ppm range — they are normally put into concentrated nutrient solutions in small proportions and included with other components such as Ca and Mg, which are present in concentrations much more in line with macro nutrients like N, P and K.



Simple model of the metal chelating process

The advantage of micro nutrients is that they are available cheaply and in high purities as heavy metal sulfate salts. These however have the problem of leading to relatively unstable cations in solution, making the preparation of

concentrated micro nutrient solutions with pure sulfates impractical (unless you want to see how a gallon of rust looks like). However we can chelate the cations as they come out of these sulfates, using a chelating agent, in order to prevent any precipitation issues. In this article I am going to walk you through the preparation of a DIY chelated micronutrient concentrated solution. This is much cheaper than buying the heavy metal chelates, which can be 3+ times more expensive. To prepare this solution you'll need to buy the chemicals shown in the table below. The table includes links to buy all the different substances mentioned plus their cost (without shipping).

Link	Price USD/lb	Weight g/gal
<u>Disodium EDTA</u>	22.96	17.0600
Ferrous sulfate heptahydrate	15.99	9.4211
Zinc sulfate monohydrate	9.49	0.1039
Manganese sulfate monohydrate	14.99	1.1646
Copper sulfate pentahydrate	20.99	0.0595
<u>Sodium Molybdate</u>	19.99	0.0191
Boric acid	10.95	3.3384
Total Cost	115.39	

List of salts to prepare a DIY chelated micronutrient concentrated solution. This concentrated solution is to be used at 5mL per liter of final feeding solution.

In order to prepare the solution you also need a scale that can weight with a precision of +/- 0.001g (this is my low cost recommendation) and a container where you can store 1 gallon of solution. Please note that these solutions have to be prepared with distilled water, with RO water you might still run into some issues in the process. To prepare the solution carry out the following steps (the weights to be used are specified in the table above):

1. Wash your container thoroughly with a small amount of

distilled water

- 2. Fill your container with half its volume of warm distilled water (30C, 86F)
- 3. Weight and add the disodium EDTA, stir until it is completely dissolved (this can take a while).
- 4. Weight and add all the remaining micro nutrients one by one in the order given above, stirring till each one is fully dissolved before adding the next.
- 5. Fill the container to its final volume using warm distilled water.
- 6. Let the solution cool before closing the container.
- 7. For longer half-life transfer to a container that is opaque to UV light.

This solution is prepared to give you the heavy metal concentrations of the Hoagland nutrient solution (a very common set of ratios used in scientific research for growing plants) when used at a ratio of 5mL per every liter of final feeding solution (18.92mL per gallon). The links given above are for 1lb of each product, with this you should be able to prepare at least 53 gallons of the concentrate, which will allow you to prepare 10,600 gallons of final feeding solution. The first salt you will run out of is Fe, but some are used so sparingly that you should be able to use them for the rest of your life without needing to buy any more (like copper sulfate and sodium molybdate). For less than 120 USD you will be able to have enough solution for probably the rest of your life — if you're a hydroponics aficionado — or even an entire crop cycle if you're a commercial grower.

This preparation is not without problems though, since the chelates are all prepared in situ they will take a substantial amount of time to reach their thermodynamic equilibrium, meaning that it cannot be used to soon or some of the metals might not be fully chelated. To obtain the full metal chelating effect an excess of around 25% of disodium EDTA is also used, which means that this micro nutrient solution

contains more free EDTA than a solution prepared with the chelates. Another issue is that all heavy metals are chelated with EDTA, which might not be optimal depending on your growing conditions. The EDTA chelates are also less stable against UV light and are also more easily attacked by oxidants. Another final issue is that the solution above contains no preservatives and fungi generally like to feast on this sort of micronutrient containing solutions. It is therefore reasonable to avoid preparing any large amounts of the above, as a solution prepared as instructed is normally expected to spoil in 3-4 weeks.

With this in mind, the above is not a perfect but a low cost and practical solution for those who want to start preparing their own nutrient solutions and avoid paying the high prices of some commercial nutrients just because of their micro nutrient contents. The above gives you a versatile micro nutrient concentrate that is bound to be adequate for growing almost all plants.

## How to prepare pH 4 and 7 buffers from scratch without using a pH meter

I wrote a post in the past about how you could prepare pH buffers in order to calibrate your pH meter if you happen to already have a calibrated pH probe. This can generate decent results if the initial calibration of the probe is excellent and the sensitivity of the probe is high. This however might not be a possibility for some people — given that their pH probe might not be calibrated to start with — so in today's

post I am going to tell you how you can prepare your own pH 4 and 7 buffers without having any other tools but a scale, distilled water and some raw salts. This tutorial will be made assuming you're preparing 500mL of each buffer but feel free to scale this up or down as you wish (these buffers are meant to give you a total 0.1M buffer concentration). Note that pH depends on temperature, these buffers are meant to give pH values of 4 and 7 at 25C.



To prepare these buffers you will need the following materials:

- A scale that can weight with a precision of +/- 0.001g
- Potassium citrate (food grade)
- Anhydrous Citric acid (food grade)
- Potassium monobasic phosphate (food grade)
- Potassium dibasic phosphate (food grade)
- Distilled water
- Two clean glass bottles to prepare and store the buffers. (I would recommend these, but any clean glass containers would do)

Follow these steps to prepare the pH 4 buffer:

- 1. Weight **exactly** 5.259g of potassium citrate and transfer that amount to the glass bottle
- 2. Weight **exactly** 6.309g of citric acid and transfer the solid to the same glass bottle

- 3. Fill the bottle to around 250mL using distilled water
- 4. Mix the solids using a glass rod or any other inert mixing utensil until fully dissolved
- 5. Fill the bottle to 500mL using distilled water.
- 6. Label the flask clearly so that you know this is the pH 4 buffer

Follow these steps to prepare the pH 7 buffer:

- 1. Weight **exactly** 3.369g of potassium dibasic phosphate and transfer that amount to the second glass bottle
- 2. Weight **exactly** 4.172g of potassium monobasic phosphate and transfer the solid to the same glass bottle
- 3. Fill the bottle to around 250mL using distilled water
- 4. Mix the solids using a glass rod or any other inert mixing utensil until fully dissolved
- 5. Fill the bottle to 500mL using distilled water.
- 6. Label the flask clearly so that you know this is the pH 7 buffer

The above should provide you with pH 4 and 7 buffer solutions that should be relatively precise. The exact volume of the solution is not critical, as the volume only has a strong effect on the buffering capacity but not on the final pH, especially at relatively high buffering strengths. However, if you want to have more precision use 500mL volumetric flasks to prepare the solutions. The error in these buffers will depend on the purity of the salts used — which is why higher purity food grade salts are recommended above — as well as in the accuracy of the weighting and transferring processes. In order to obtain a higher accuracy you would need to purchase more expensive analytical grade salts and also use volumetric flasks to prepare the solutions, so that you can prepare them at the exact concentration intended.

Another limitation of the above buffers is that they do not contain any sort of preservative and they are both prepared with food grade substances that can attract fungi and bacteria. For this reason the above buffers will probably not last for a significant amount of time and should probably be discarded within a couple of weeks. However the chemicals used here are very cheap so — with the amounts purchased above — you should be able to prepare as much buffering solution as you might need. Note that the solutions can also be frozen in order to increase their shelf life, although keep in mind that since pH depends on temperature you will need to wait for them to reach room temperature before taking a reading.

It is also worth mentioning that these buffers will both be completely transparent, since they are prepared without any dies in order to give the maximum possible accuracy in the pH. However you can add a very small amount of food coloring to each one to provide a distinct color without causing a significant change in the pH, less than half a drop should be enough to give your solutions a distinct hue.

I would advice you do a pH check with a pH meter calibrated using a normal commercial solution the first time you prepare these solutions. This is just to be sure that you followed the procedure correctly and the resulting buffer is of the intended quality. Once you do this you should be able to create as much buffer as you desire without any problems. Leave a comment with your experience!

## Why TDS is NOT equal to Total Dissolved Solids in hydroponics

Electrical conductivity is a very commonly used measurement in hydroponics, yet a very poorly understood one. I have written

several posts about conductivity in the past (1,2,3) and today I want to talk about the use of the term "Total Dissolved Solids" and the poor usage of the unit "ppm" in order to express a measurement of electrical conductivity. In this article I will walk you through why this term exists in the first place and why its use in hydroponics is terribly misleading for growers.

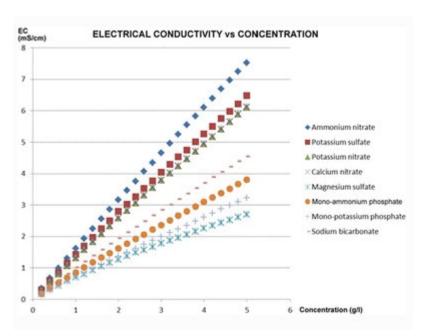
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Conductivity as a function of NaCl concentration (taken from <a href="here">here</a>)

Conductivity is just a measure of how easy it is for an electrical charge to go from one electrode of a certain area to another. It's generally expressed in mS/cm, which is a measurement of conductance (the opposite of resistance) and area (the area of the electrode). How in the world do we get from this to a measurement like "ppm", which measures the concentration of something in mg/L? What does a measurement of 500 ppm even mean? What is it that we are expressing a concentration of?

The answer lies in the practical uses of conductivity and a simplification to make the evaluation of water sources easier. Conductivity is generally linearly proportional to the amount of a pure salt dissolved in solution at low concentrations. For a pure salt like table salt (NaCl) the higher the concentration of the salt in solution the higher the conductivity (you can see this in the image above). People working on water quality realized that they generally dealt with similar salt combinations (Mg and Ca carbonates and possibly some Na and K chlorides) so they decided to use some standard salt mixtures (say KCl, NaCl or some mixture of Ca/Mg/K/Na salts) and then use conductivity as a proxy for the concentration of these things that are actually in solution. So the "ppm" that your EC meter reads is just the equivalent conductivity of some standard. A meter reading 500 ppm in conductivity is telling you "your solution has the same

conductivity as a solution of the standard at 500 ppm". The "standard" can change — as mentioned before — which is why there are several different TDS scales. One meter might be telling you it's the same conductivity as a solution of KCl with that concentration, while another might be in NaCl.



Conductivity curves of different salts used in hydroponics (taken from <a href="mailto:this.article">this.article</a>)

The above is very useful when you're measuring things that tend to be similar but this becomes a complete nightmare when the composition of what you're measuring can change substantially. In hydroponics you have a wide variety of different salts, all with very different conductivity values at different concentrations. Look at the graph above, which shows the conductivity as a function of concentration for 8 different salts commonly used in hydroponic culture. If you prepare three solutions, one with 1000 ppm solution of potassium sulfate, another with 1000 ppm of monopotassium phosphate and another with 1000 ppm of ammonium nitrate and measure them with your conductivity meter they would all give very different results. The meter might be close to 0.95mS/cm for the monopotassium phosphate, but it might read almost 1.5 mS/cm for the potassium sulfate. Both solutions have 1000 ppm of "total dissolved solids" but the conductivity meter is telling you one has 500 ppm and the other almost 800 ppm, none

of them even close. This is because "total dissolved solids", as used in water quality measurements, is a meaningless measurement in hydroponics as it relates to the actual ppm values of things dissolved.

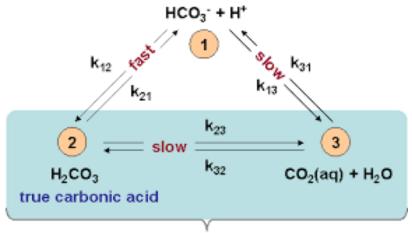
This is the main reason why you should never compare the EC values of nutrients that contain different ratios of salts, because they are simply not the same. One nutrient might give you 100 ppm of potassium at some EC level, while another might give you 200 ppm. Thinking that having the same EC level means that both are at the same "strength" is a big mistake, since this is never going to be the case when two nutrient solutions are mixed with different ratios of nutrients. This is also why comparing vegetative and bloom formulation EC values is not correct. A solution in veg might contain a lot more of nitrates while a solution in bloom might contain more phosphates. As we saw above this might mean that a solution of the "same strength" might actually have a significantly lower measured EC value.

Since the TDS measurement is not telling you anything about "total dissolved solids" in hydroponics, you should avoid using it to avoid confusion. This is important since nutrient concentrations are usually expressed in ppm as well, ppm of actual nutrients dissolved in solutions. Instead use the normal conductivity measurements of your meter in conductance per area. You should also take care to only use EC values to talk about comparative strength when you're talking about a formulation where the ratios of nutrients remain the same. If that's not the case, then you should not talk in comparative terms between the two solutions as this might deviate a lot from reality.

My advice is to not think in EC terms to begin with, but to think about nutrient concentrations, prepare solutions that match the concentrations you want and then use the EC of those solutions as references to know whether they are prepared correctly or not. The conductivity should be a measurement used for confirmation but not as a guiding principle. For example the aim should be to "prepare a solution containing 150 ppm of N and an K:N ratio of 1.2" not to "prepare a solution with an EC of 1.2 mS/cm".

# Understanding the carbonic acid/bicarbonate buffer in hydroponics

I have written several articles before about pH and it's importance in hydroponic culture (1, 2, 3, 4). However I have yet to write a detailed explanation of one of the most important buffering systems in hydroponics, which is the carbonic acid/bicarbonate buffer. This buffer is significantly more complicated than the simpler buffer created using phosphoric acid species, as it not only relies on ions present in solution but also on the partial pressure of carbon dioxide in the atmosphere. In this article I will attempt to explain this buffering system in detail, shining some light into the limitations of this buffer and how changing different key variables can fundamentally affect the way it works in hydroponics.



composite carbonic acid

Chemical reactions involved in the carbonic acid/bicarbonate buffer. Taken from <a href="here">here</a>.

A buffer is nothing more than a pair of chemical species in solution that are present at a certain pH, that can react with additional  $H_3O+$  or OH- ions that are introduced into the solution. Since these ions control the value of pH, anything that prevents their concentration from changing will keep the pH stable. Distilled water, for example, has absolutely no buffering capacity since within it there is nothing that can react with incoming  $H_3O^+$  or  $OH^-$  ions that are added to the solution. Distilled water should therefore have a pH of 7.0, it does not because we live in an environment where an acid can always be generated from the air. This acid — carbonic acid — is generated in water whenever it's put into contact with a carbon dioxide containing atmosphere. This makes distilled water have a pH of around 5.6.

To be able to calculate the pH we need to consider all the chemical equilibrium reactions that happen, these are summarized here and in the image above. We must consider that carbon dioxide will dissolve in water to always satisfy Henry's law, that dissolved carbon dioxide will be in equilibrium with carbonic acid, that carbonic acid can dissociate into a  $\rm H_30^+$  ion and a bicarbonate ion and that a bicarbonate ion can further dissociate into an additional  $\rm H_30^+$  ion and a carbonate ion. To solve all of this we must also

consider that charge neutrality must be preserved, meaning that the sum of all molar charges of all positive ions must be equal to the molar charges of all negative ions. To carry out these calculations I routinely use the freely available <a href="Maxima software">Maxima</a>. Below you can see the code I use to solve this system in Maxima (constants are taken from <a href="here">here</a>):

```
[kw : 10^{(-14.0)}, kh: 1.7*10^{(-3.0)}, kc1: 2.5*10^{(-4.0)}, kc2: 4.69*10^{(-11.0)}, co2: 1.32*10^{-5}]; log10(x) := log(x)/log(10) ; pH(x) := float(-log10(x)); float(solve([h*oh=kw, h = 2*co3+hco3+oh, kh=h2co3/co2, kc1=(hco3*h)/h2co3, kc2=(co3*h)/hco3],[oh, co3, hco3, h2co3, h]));
```

This is the solution obtained for the molar concentrations (rounded for clarity):

```
oh = 4.21*10^{-9}

co3 = 4.68*10^{-11}

hco3 = 2.36*10^{-6}

h2co3 = 2.24*10^{-8}

h = 2.37*10^{-6}
```

After executing this code you will get several different possible solutions, but the only one that interests us is the one where the  $H_30^+$  (h) concentration is a positive number (this solution is showed above). We can then use the pH function to calculate the value of pH for this  $H_30+$  concentration, which gives us a value of 5.62, this matches the real measurement of a distilled water solution at 25C under a 387ppm carbon dioxide atmosphere. Note that the amount of none dissociated acid in solution is very small. Taken to mass, the concentration of carbonic acid is 0.00138 ppm. However the concentration of bicarbonate is significantly greater, at 3.6 times the concentration of undissociated carbonic acid. This

explains why the pH drops so much, since a significant amount of the generated carbonic acid ends up dissociating and contributing  $H_3O^+$  ions to the solution. This also shows you how little acid is needed to drop the pH of an unbuffered solution.

To create the buffer with the biggest possible strength we would need to add enough strong base to shift the pH to the point where the pH equals the pKa (which is just - Log(equilibrium constant)) of the joint reactions created from the reaction of carbon dioxide with water to create carbonic acid and the subsequent dissociation of this acid into bicarbonate and  $H_3O^+$ . This point is at 6.3 under atmospheric conditions at 25C. This can be achieved with the code below:

```
[kw : 10^{(-14.0)}, kh: 1.7*10^{(-3.0)}, kc1: 2.5*10^{(-4.0)}, kc2: 4.6910^{(-11.0)}, co2: 1.32*10^{-5}, h:10^{(-6.3)}; float(solve([hoh=kw, base+h = 2co3+hco3+oh, kh=h2co3/co2, kc1=(hco3h)/h2co3, kc2=(co3h)/hco3],[oh, co3, hco3, h2co3, base]));
```

This is the solution obtained for the molar concentrations(rounded for clarity):

```
oh = 1.99*10^{-8}

co3 = 1.04*10^{-9}

hco3 = 1.11*10^{-5}

h2co3 = 2.24*10^{-8}

base = 1.07*10^{-5}
```

The pH here is set to 6.3 and we can see that to get there we would need to add a base at a concentration of  $1.07*10^{-5.0}$ . If this base was KOH this would imply adding it at a rate of 0.6 ppm. We can see how the pH changes as a function of adding base or acid from this point. If at this point we decided to double the addition of strong base we would get to 6.57, tripling it would take us to 6.73 and adding 10 times more base would take us to 7.25. The buffer is indeed resisting the

increase in pH by basically drawing  $\mathrm{CO}_2$  from the air to react with the incoming base as base is added to the solution. However you might notice that under equilibrium conditions the buffering capacity of this system is very low. Just 6 ppm of a KOH equivalent strong base addition can strongly affect the pH — taking it from 5.6 to 7.25 — so how can the carbonic acid/bicarbonate buffer be effective at all in hydroponics?

The answer is in the first image in this post. The equilibrium reaction between carbonic acid and water plus carbon dioxide in water (k23/k32) is fundamentally slow. We can take advantage of this by generating larger amounts of carbonate species in solution through the use of exogenous carbonate or bicarbonate additions and then setting the pH at a lower value to generate more carbonic acid, this acid will then take some significant time to reach equilibrium. This is the reason why using tap water with a significantly high alkalinity can provide a surprisingly stronger buffer than what would be expected at equilibrium and it also has some interesting consequences in the use of nutrient solutions.

Let's consider a case where there is no decomposition of carbonic acid — let's suppose it's extremely slow — and say we add 100 ppm of potassium carbonate into a solution and then set the pH back to 5.8 using phosphoric acid. In this case the predominant reactions in solution would be the dissociation of dihydrogen phosphate to hydrogen phosphate and  $H_3O+$  and the carbonic acid dissociation discussed before. In order to properly consider this case we must also introduce two additional equations, mainly the mass balance equations for the phosphate and carbonate species, since this time we are assuming no carbon dioxide is ever lost to the atmosphere. Note that I have changed the equilibrium constant for the carbonic acid reaction here to  $10^{-6.3}$  where carbonic acid is now "apparent carbonic acid". You can see the equation system and solution below:

[kw :  $10^{(-14.0)}$ , kh:  $1.7*10^{(-3.0)}$ , kc1:  $10^{-6.3}$ , co2:  $1.32*10^{-5}$ , kp: $10^{-7.2}$ , total\_p:  $1.7*7.2310^{-4}$ , total\_c:  $7.23*10^{-4}$ ];

float(solve([h\*oh=kw, total\_c=hco3+h2co3, total\_p=h2po4+hpo4,  $2*total_c+h = hco3+oh+h2po4+2*hpo4, kc1=(h*co3h)/h2co3, kp=(hpo4*h)/h2po4],[hco3, h2co3, h2po4, hpo4, h, oh]));$ 

This is the solution obtained for the molar concentrations(rounded for clarity):

hco3 =  $1.72*10^{-4}$  h2co3 =  $5.50*10^{-4}$ h2po4 = 0.00118hpo4 =  $4.64*10^{-5}$ h =  $1.60*10^{-6}$ 

The final pH of this solution is very close to 5.8 and the concentration of P is 47.9 ppm with K at 38.10 ppm. Notice however that apparent carbonic acid has a concentration of 5.50\*10<sup>-4</sup> M, which implies that the system is not equilibrium since this amount is significantly larger than what we would expect from Henry's law. If we reduce the concentration of carbonic acid to half then the pH will increase to 6.01, as we would expect from extracting an acid from the solution. The implication is that — with time — the pH of this solution is going to slowly increase, as carbonic acid decomposes and the solution reaches an equilibrium with the atmospheric carbon dioxide level. This is also why nutrient solutions that are prepared with tap water high in carbonates and then aerated will tend to show a rapid increase in pH — even if the solution is not fed to plants — as the reaching of equilibrium is accelerated by the agitation of the solution and the contact with air (that allows CO<sub>2</sub> in solution to escape).

As soon as the above solution is prepared it offers a substantially superior buffering capacity when compared with a solution containing only phosphates. This is why water with high alkalinity tends to provide better pH stability in drain to waste type systems when compared with solutions prepared with RO water. This water contains a significant amount of carbonates that are turned into carbonic acid and bicarbonate as soon as the pH is lowered to the pH range used in hydroponics. As long as the solution is used quicker than the carbonic acid decomposes, there will be a substantial increase in pH stability.

If you are using RO water or water with low alkalinity to prepare your solutions you can obtain a similar effect by adding 100-200 ppm of potassium carbonate before you start preparing the nutrient solution, you can similarly use bicarbonate but I would recommend using potassium carbonate, as it is cheaper. It would also be advisable to use the solution as fast as possible, since time will cause the solution to reach equilibrium and the pH to increase. This effect will take much longer if the  $\mathrm{CO}_2$  concentration is higher — which is true for setups that use enriched  $\mathrm{CO}_2$  — or if the temperature is lower, which increases the solubility of  $\mathrm{CO}_2$ .

#### Hydroponics nutrients and microgreens

One of the most important goals in microgreens is to maximize the amount of weight gained by shoots from seed to harvest. Since the entire upper body of the plant is harvested and plants are sold by weight, maximizing the weight gain is vital in order to obtain the highest possible margins in a crop cycle. Hydroponically cultured microgreens offer the grower an unprecedented control over the microgreens' nutrition, with the ability to tightly control nutritional parameters in order

to maximize this weight. In this article we are going to take a look into the scientific literature surrounding microgreens and what we know about maximizing their yield and quality using nutrient solutions. I will use the table below to reference different articles in the literature.

Number	Species	Studied	Link
1	Broccoli	FlograGro, sterile, compost	https://www.frontiersin.org/articles/10.3389/fnut.2017.00007/full
2	Purple Cabbage	Nutrient sol conc	https://www.scielo.br/scielo.php?pid=S1983-21252019000400976&script=sci_arttext&tlng=en
3	Table Beet	Calcium Nitrate	https://www.tandfonline.com/doi/abs/10.1080/19315261003648241
4	Radish	Calcium Chloride	https://bearworks.missouristate.edu/theses/3328/
5	Basil	Sodium Selenate	https://onlinelibrary.wiley.com/doi/abs/10.1002/jsfa.9826

Published articles talking about hydroponic nutrients and microgreen yield or quality

Despite the overwhelming growth in the microgreen industry during the past 10 years, the amount of research looking into microgreen nutrition has been surprisingly limited, with only a handful of papers looking at the relationship between nutrition and yields or quality. Paper one contains a comparison between microgreens grown in either compost, sterile water or a solution using a 0.4% FloraGro Advanced Nutrient solution (4mL/L). The results show clear weight benefits from using hydroponic nutrients, with the weight being markedly higher (mean of 24.64g vs 21.01g) between the sterile and hydroponic treatments. However the concentration of different minerals was actually lowest in the plants using a hydroponic nutrient.

 
 Table 3: Shoot fresh weight per plant and per population of `Early Wonder Tall
 Crop' table beet at 15 days after planting in response to seed ball and fertilizer treatments.

Seed treatment (ST)	Fertilization treatment (FT) <sup>a</sup>	Shoot fresh weight (g·m <sup>-2</sup> )	Shoot fresh weight (mg/shoot)
Nontreated	Check CN SF 0.5 CN + 0.5 SF	2705h 4207f 4676e 4676e	166e <sup>b</sup> 268d 268d 284cd
Pregerminated <sup>c</sup>	CN + SF Check CN SF 0.5 CN + 0.5 SF CN + SF	5469cd 3391g 5087de 5649c 6294b 6937g	299cd 298cd 362b 306c 362b 415a
Interaction LSD <sub>0.05</sub> Significance	3.11.5	444	36
ST FT		***	***
ST × FT		*	**

#### Table taken from article number three

Papers three and four look at different forms of Ca nutrition - either Ca chloride or Ca nitrate - and different ways to apply this treatment to see if it makes a difference in microgreen production. Paper three, shows a statistically significant gain in weight when using calcium nitrate, either applied into the media pre-cultivation or applied within a nutrient solution. The best results were found when both treatments were carried out and represented an increase of more than double in terms of weight over the control. The fact that paper four fails to show a consistent increase in yields using Ca chloride, suggests that this has to do mainly with the nutritional contribution of the nitrate and not the calcium ions.

Paper two is rather interesting, as it looks into different nutrient solution strengths (either 0, 50 or 100%) using a solution published for hydroponic forage. The results — in the table below — clearly show that there is a strong weight gain as the nutrient solution concentration increases, again showing that at a full strength solution there is an expected increase of more than 2x in the final weight. However this comes — in agreement with paper one — at the potential expense of nutritional value. The paper shows a significant decrease

<sup>\*,\*\*, \*\*\*</sup>Significant at P = 0.05, P = 0.01, or P = 0.001, respectively. 

\*\*Check = no additional fertilizer added to the peat-lite; CN = preplant addition of solid  $Ca(NO_3)_2$  at 2000 mg·L<sup>-1</sup> of N (150 mL·L<sup>-1</sup> of peat-lite); SF = daily application of 150 mg·L<sup>-1</sup> of N solution fertilizer from 21-5-20; 0.5 = one-half rate.

\*\*Means within a column followed by the same letter are not significantly different by LSD<sub>0.05</sub>.

\*\*CPregerminated = seed ball incubation in fine (grade 5) vermiculite with 150% water for 5

in carotenoid concentration when nutrient solution strength increases, which the paper hypothesis is caused by high nutrient concentrations slowing down plant metabolism. This hypothesis is however hard to reconcile with the larger and heavier plants.

Table 2 Effect of cultivation substrate and nutrient solution concentration on shoot fresh matter yield (SFMY) and shoot dry matter yield (SDMY), shoot height at harvest (SHH), and total soluble solids (TSS) content of purple cabbage microgreens (*Brassica oleracea* var. capitata f. rubra) in a recirculating system. Porto Alegre, Brazil, 2018.

		SFMY	SDMY		TSS
Factors	Trataments —	(g m <sup>-2</sup> )		SHH (cm)	(°brix)
	S1 (CSC <sup>®</sup> Vermiculite)	1829.52 <sup>ns</sup>	74.01 <sup>ns</sup>	6.86 <sup>ns</sup>	3.62 <sup>ns</sup>
Substratos	S2 (Beifiur <sup>®</sup> S10)	1761.30	79.21	6.89	3.64
	S3 (Carolina Soil® seedling)	1795.73	75.70	6.83	3.73
	S4 (Carolina Soil® organic)	1793.57	81.10	7.19	3.82
	C1 (0%)	1111.31 c*	64.34 c	5.00 c	5.03 a
Concentration Nutritive Solution **	C2 (50%)	1933.16 b	78.62 b	7.36 b	3.37 b
6.963	C3 (100%)	2340.62 a	89.59 a	8.48 a	2.72 c
	Mean	1795.03	77.52	6.94	3.70
	CV (%)	16.34	11.29	6.83	14.45

nsnot significant by the Tukey's test at 5% probability;

#### Table taken from article number two

Article five is also an interesting example of the use of microgreens to carry out antioxidant supplementation. Sodium selenate was used to prepare a solution to treat basil seeds and the resulting microgreens were found to be fortified with selenium. This might be an interesting way to incorporate mineral micro nutrients into microgreens and therefore increase their presence within our diet. However there is also the potential to hyperaccumulate these nutrients, so experiments of this kind should not be done with adequate care and lab analysis to ensure proper doses of these micro nutrients.

From all of the above it seems quite clear that the research of hydroponic nutrients in microgreen production is in its very early infancy. So far only a handful of research papers have been published on the subject and the conclusions so far seem to be that hydroponic nutrient solutions — in a couple of different forms — tend to significantly increase microgreen

<sup>\*</sup>means followed by the same letter in the columns are not different by the Tukey's test at 5% probability.

<sup>\*\*</sup>initial electrical conductivity (ECi) established for the three evaluated nutrient concentrations (0%, 50%, and 100%) in the nutritive solution: C1 = 0.0; C2 = 1.20, and C3 = 2.0 dS  $m^{-1}$ , respectively.

production weights. However it is also clear that there is a strong interaction with the nutritional value of the microgreens and using nutrients can in fact lead to decreases in the nutritional value, despite the significant weight gain from the process.

The echoes of the above can be seen in a wide variety of anecdotal experiences on youtube channels and forums. Growers running side by side experiments seem to have found the same phenomena (see <a href="this video">this video</a> for an example), where adding nutrients increases yields significantly but at the expense of some of the flavor — and potentially nutritional — characteristics of the microgreens. Some growers have therefore chosen to avoid nutrients — to preserve flavor qualities — while others have chosen to use nutrients because of the increases in marketable appearance and yield.

There is a lot of research to be done on the subject. It would certainly be interesting to find out if we could somehow have the best of both worlds.