A TALE OF TWO IONS: AN ANALYSIS OF CALCIUM AND CHLORIDE IONS IN WATER SOURCES FROM MADISON, NEW JERSEY

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ABSTRACT

Understanding the quality of our local water is becoming a paramount issue for New Jersey communities. In this study, the team investigated the contents of different samples of water including the Borough of Madison's tap water, water from varying nearby locations, and Stop-and-Shop spring water. The team focused primarily on the calcium and chloride ions in the samples through titrations, and suggested possible sources of these ions. A qualitative study using the persistent presence of soap bubbles explored the hardness of several water samples. Conductivity and pH tests provided insight into the total ion concentration and the solubility of calcium carbonate. Upon examination of the water samples, Madison tap water was found to have exhibited higher concentrations of calcium and chloride ions than spring water. Furthermore, the team tested samples of water from various streams that feed into the Great Swamp aquifer; this allowed us to track the presence of ions from the source. This thorough examination of water analyzes the ions present in different types of water and offers possible sources that the contaminants could have originated from.

INTRODUCTION

Water is a very important compound. This simple substance, made of only one oxygen atom and two hydrogen atoms, is inescapably essential to life. Comprising approximately 60% of the human body, one cannot survive for more than three days without this vital nourishment [1]. Water plays a prominent part in the body: it regulates body temperature, transports nutrients and wastes (most of our bodily fluids are based on water), and maintains homeostasis. Due to its importance, water is one of the key components in determining quality of life. As a result, the quality of drinking water is a cause of major concern. Although nearly 70% of the Earth is composed of water, only 1% is available for drinking purposes [2].

Known as the "universal solvent," it is nearly impossible to obtain pure water devoid of all impurities [3]. Today, water must be purified before consumption in order to mitigate the effects of harmful substances. Most impurities are removed from the drinking water through the processes of screening, sedimentation, filtration, chlorination, and irradiation. Nonetheless, even with this intense processing technology, some contaminants will still remain.

Two of the most common ions found in drinking water are calcium and chloride. When calcium, magnesium, or iron ions are present in drinking water in high concentrations, the water is considered "hard." Although soap cannot form lather as readily in the presence of calcium ions, hard water is not particularly harmful. There are no known health effects associated with

calcium and magnesium in water. However, mineral buildup caused by hard water affects water heating systems. People use many techniques to counter hard water, including boiling, treating water with sodium carbonate and lime, and filtrating through natural and artificial zeolites.

There are many reasons why calcium and chloride ions are prevalent in water. Small amounts of calcium are found in soil deposits in the form of calcium carbonate, or limestone. Rainwater dissolves calcium carbonate and carries the dissociated calcium ions into bodies of water. The most probable explanation for the presence of chlorine ions in water is the chloride compounds used by municipalities to treat and disinfect water. Another possible source of the chloride ions is road salt, which is originally used to prevent ice.
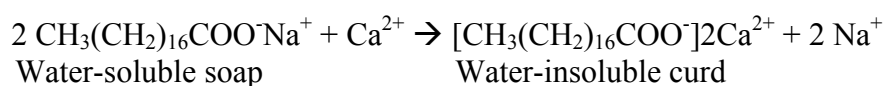
Calcium is extremely important to the human body. The heart, nerves, bones, and muscles all require calcium to function properly, maintaining a fragile balance of the internal systems. It is also needed in blood coagulation and may act as an enzyme cofactor. Calcium deficiencies can trigger side effects such as osteoporosis [4]. Similarly, trace amounts of chloride are also imperative to human health. Chloride is a key component of bile, necessary for the formation of hydrochloric acid used in the digestive system.

The ion content of Madison's tap water can be traced along the streams leading to the Great Swamp. Loantaka Brook connects with various ponds that contribute to its total water source as it continues its path toward the Great Swamp. Among those bodies of water are the collected Pond 1 and Pond 2 water samples. In terms of location, Pond 2 is found to lay upstream, north-north-east of Pond 1 and will be the first to pour into the Loantaka Brook. Another brook that will flow into the Great Swamp is the Black Brook. Water samples for this experiment were collected directly from the brook itself. Located just under the swampland, the Buried Valley Aquifer is the immediate source of Madison tap water. Alternatively, spring water can be purchased in stores; the Stop-and-Shop spring water we examined lists its water source as a Massachusetts spring. These water samples serve as the focus of this analysis.

METHODS

Water Hardness

The defining characteristic of hard water is the presence of calcium, magnesium, or iron ions. Water hardness impacts the effectiveness of soaps and detergents. This relationship is dictated by the structure and function of soap. Soap molecules are comprised of long carbon chains with two distinct ends. They contain extensive tails of non-polar, hydrophobic fatty acids, offset by a hydrophilic head made of a carboxyl group and either a sodium or potassium ion. When soaps and detergents are placed in the presence of hard water, the aforementioned "hard" ions tend to replace the sodium or potassium ions at the head of the soap molecule, forming insoluble carboxylate salts according to the reaction:

$$2\ CH_3(CH_2)_{16}COO^-Na^+ + Ca^{2+} \rightarrow [CH_3(CH_2)_{16}COO^-]2Ca^{2+} + 2\ Na^+$$

Water-soluble soap   Water-insoluble curd

To determine the relative water hardness of de-ionized water, spring water, and Madison tap water, we obtained 50 ml of each water sample and proceeded to titrate it with 1 ml increments of soap solution. After each soap increment was added, the solutions were stirred and monitored for the persistent presence of soap bubbles. If the bubbles remained for more than ten seconds, it was assumed that the titration had reached its endpoint. Taking the five distinct sets of data, we took the arithmetic mean to give us an estimate of the water hardness of the three samples.

Total Ion Concentration

Conductivity measures the ability of an aqueous solution to transfer an electrical current from one point to another. When an electric current is applied, (assuming the solution contains both positive and negative ions) the cations move toward the negatively charged cathode, while anions move to the positively charged anode. As the ions move toward the electrode of opposite charge, it is able to measure the electrolytic conduction of the solution. Several factors determine the quality of the solution's ability to conduct electricity: the concentration of ions, the mobility of those ions, the oxidation state (by way of valence electrons) and the temperature of the water. Therefore we can use conductivity as a means to calculate total ion concentrations.
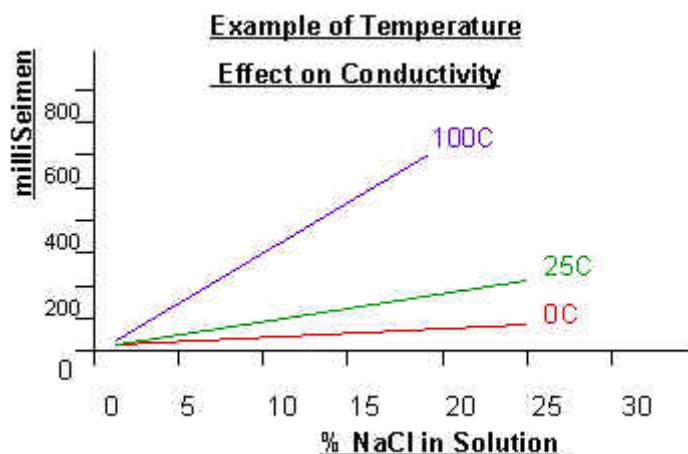

Figure 1: Graph of conductivity as a function of % NaCl in solution at varying temperatures

Temperature has a significant effect on the conductivity of solutions. Figure 1 shows that the same solution concentration at different temperatures has a highly disparate conductivity measurement [5].

To determine the conductivity of a solution, a self-contained conductivity electrode probe was used. The data was collected in a computer program named Logger Pro. Before beginning experimentation, the conductivity probe was calibrated using de-ionized water as a constant (set equal to zero) and a standard solution with a known ion dissociation rate. Afterwards, the measured conductivities of the samples (Madison tap water and spring water) were compared to the previously calibrated substances to determine a relative conductivity.

To derive an equation relating conductivity and ion concentration, we tested the conductivity of NaCl, $CaCl_2$, $Al(NO_3)_3$, and CH_3COOH solutions of known ion molarity (see Table 4 of Appendix A). Because of the direct relationship between ion molarity and conductivity, we were able to plot these four values and compute the regression curve, producing the linear graph of conductivity and ion molarity (see Figure 2). The graph provides us with a means of determining the unknown ion molarity of any water sample with measured conductivity.
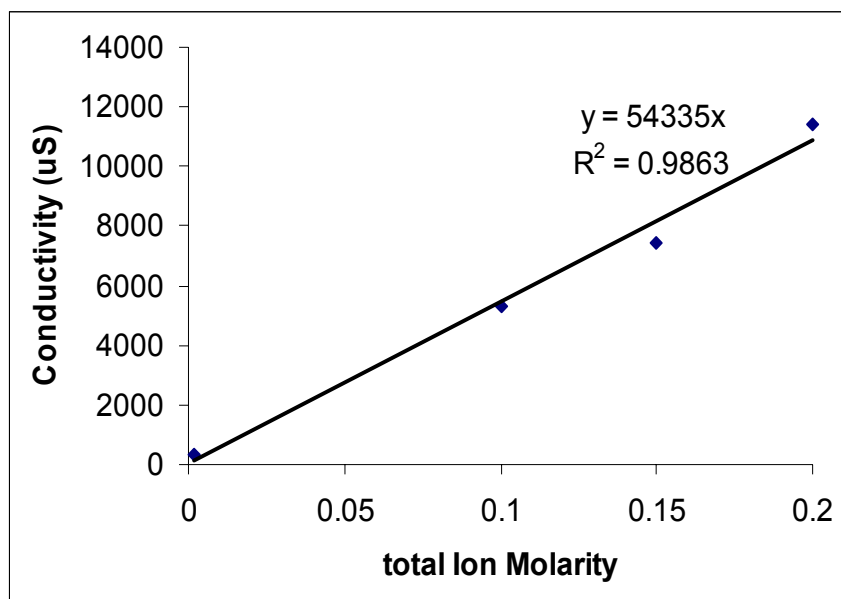

Figure 2: Observed relationship between conductivity and total ion molarity

Total Dissolved Substances

To determine the total mass of dissolved substances in the de-ionized water, spring water, Madison tap water, Pond 1 and Pond 2, we evaporated 200 ml of each sample. Any remaining solid material was assumed to have been the total dissolved solids. By finding the mass of the dried solid and comparing it to the de-ionized water sample (the control), we were able to find the arithmetic mean mass of the substances in each respective sample.

pH

pH is the measure of the acidic or basic (alkaline) nature of a solution. It is measured by the concentration of hydrogen ions present. Mathematically, it is expressed as:

$$pH = - \log [H^+]$$

A logarithmic function in base 10, the solution changes ten times in hydrogen ion concentration with every whole number increase or decrease in pH.

Most bodies of water may have varying pH readings due the substances present in the water and the bottom soil layer. Calcium carbonate present in rocks, soil, and other sediment

matter dissociate into calcium ions, which may significantly raise the pH levels and cause the groundwater to be less acidic. Conversely, rainwater may have a lower pH due to acid precipitation.

To determine the pH of a solution, an electrometric measurement was made by using a pH probe. The meter was calibrated with separate solutions of known pH values. Once the probe was calibrated, Logger Pro gathered the raw data and calculated the pH relative to the original calibrations.

Calcium and Chloride Ion Titrations

In order to test for the concentrations of Ca^{2+} and Cl^- ions within the different types of water, we used a titration process. To determine the total chloride ion concentration, we used a 0.1 M silver nitrate ($AgNO_3$) titrating solution and a Potassium chromate indicator (K_2CrO_4). When combining a solution of $AgNO_3$ with chloride ions, the following chemical reaction occurs:

$$Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$$
White precipitate
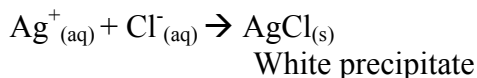
Once the Ag^+ has reacted with all of the Cl^- present in the water, the excess silver ions react with the chromate ion, and the reaction is deemed to have reached its endpoint. The excess silver ions react in the following manner:

$$Ag^+_{(aq)} + CrO_4^-{}_{(aq)} \rightarrow Ag_2CrO_{4\ (aq)}$$
Orange
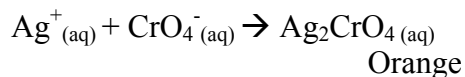
The Ag_2CrO_4 formed at the endpoint changes the water from a tinted whitish, yellow color to a bright orange hue. At the first sight of new color, exactly enough $AgNO_3$ has been used to react with all the Cl^- present in the water sample. By using a simple mathematical calculation, we were able to determine the total Cl^- ion concentration in mg/L.

L of $AgNO_3$ x Molarity of $AgNO_3$ x $\dfrac{1\ mol\ Cl^-}{1\ mol\ Ag^+}$ x $\dfrac{35.453g\ Cl^-}{1\ mol\ Cl^-}$ x $\dfrac{1000mg\ Cl^-}{1\ g\ Cl^-}$ = Cl^-ion mg/L

This same process was used in order to test for the concentration of calcium ions present in the water. In order to test for this type of ion, we made a solution of 0.02 M ethylenediaminetetracetic acid (EDTA) solution and added several drops of the Eriochrome Black T Indicator. When mixing a solution of EDTA with calcium ions, the following chemical reaction occurs:

$$EDTA^{-2} + Ca^{2+} \rightarrow Ca(EDTA)$$
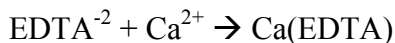
When all of the EDTA has reacted with the Ca^{2+}, the excess EDTA begins to react with the Eriochrome Black T Indicator, changing the water from a purple to a red-orange color. This point is indicative of when enough EDTA has been used to react with all of the Ca^{2+}. In a

similar manner as before, we were able to calculate the amount of calsium ions present in the water.

RESULTS

Water Hardness

Of the four samples, spring water exhibited the highest level of water hardness, requiring 49.33 ml of soap solution to reach its endpoint (see Table 1). However, this value was reduced by approximately one third (32.0 ml) when the spring water was boiled. Madison tap water (27.0 ml) had a low level of water hardness in comparison to the two spring water averages. All of our water values were compared to a standard solution of calcium acetate, which had a measured water hardness of 24.33 ml.

Table 1 Observed volumes of soap solution required for soap bubble persistence

	Calcium Acetate (mL)	Madison (mL)	Spring (mL)	Boiled Spring (mL)
Team 1 - Trial 1	30.0	38.0		
Team 1 - Trial 2	23.0	30.0	74.0	
Team 1 - Trial 3	32.0	30.0		32.0
Team 2 - Trial 1	15.0	24.0	38.0	
Team 2 - Trial 2	17.0	24.0	36.0	
Team 2 - Trial 3	29.0	16.0		
Average	**24.3**	**27.0**	**49.3**	**32.0**
St Dev	7.15	7.46	21.4	Undefined

Total Dissolved Substances

Madison tap water contained the greatest amount of total dissolved substances (0.528 g/L; see Table 2). The dried solid mass of spring water was significantly lower than that of Madison tap water; spring water contained approximately half the amount compared to tap water (0.286 g/L). After filtration, Pond 1's total dissolved substance mass was 0.399 g/L, the median value of the three samples tested.

Table 2 Total Dissolved Substances (g/L)

	Spring Water (g)	Madison water (g)	Pond 1 (g)
	0.285	0.485	0.600
	0.285	0.570	0.580
	0.240	0.505	0.135
	0.275	0.535	0.280
	0.345	0.545	
Average	**0.286**	**0.528**	**0.399**
St Dev	0.038	0.033	0.229

<u>Total Ion Concentration</u>

Conductivity tests showed that Pond 2 had the greatest total ion concentration, while the rain water had the least (see Table 3). Pond 2 had the highest concentration of ions. In contrast, rainwater contained no detectable ions. Madison tap water and spring water had very similar ion concentrations. When the samples were heated, the conductivities of both spring and Madison tap water decreased.

Table 3 Conductivities and Molarities of standard solutions and water samples

	Conductivity (μS)	*Ionic Concentration (mM)*
.05M NaCl	5338	98
.05M CaCl$_2$	7458	137
.05M CH$_3$COOH	358	6
.05M CaCO$_3$	5.0	0.08
.05M Al(NO$_3$)$_3$	11432	210
Madison	630	12
Heated Madison	604	11
Pond 1 (7/18)	874	16
Pond 1 (8/5)	667	12
Pond 2 (7/18)	2186	40
Pond 2 (8/5)	1529	28
Black Brook	459	8
Spring	656	12
Heated Spring	426	8
Round Pond	138	3
Rain	0	0

<u>pH</u>

Results from the pH tests show that Pond 2 (8/5) had the lowest pH reading (5.94; see Table 4). Spring water, Madison water, Black Brook and Pond 1 all had similar pH values, varying between 6.08 and 6.22. The two samples, with the highest pH readings, were rain and de-ionized water at pH 7.03 and 7.68, respectively.

Table 4 pH of tested water samples

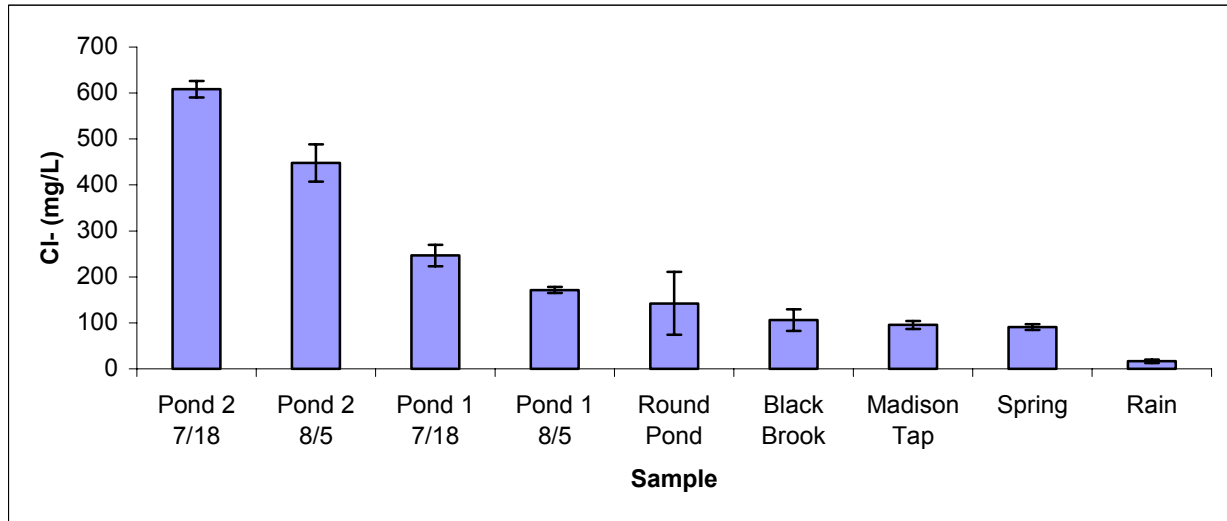|	*Trial 1*	*Trial 2*	*Trial 3*	***Average***	*St Dev*
De-ionized	7.82	6.91	8.30	**7.68**	0.71
Spring	6.20	6.11	6.36	**6.22**	0.13
Madison	6.04	6.09	6.20	**6.11**	0.08
Pond 1	6.14	6.01		**6.08**	0.09
Pond 2	6.01	5.86		**5.94**	0.11
Rain	7.65	6.41		**7.03**	0.88
Black Brook	6.21	6.16		**6.19**	0.04
Round Pond	6.51	6.28		**6.40**	0.16

Total Chloride Ion Concentration

Figure 3: Chloride ion concentration in different water samples

Figure 3 indicates that the water taken from Pond 2 on 7/18 has the greatest chloride ion concentration (608.37 mg/L), while rain water has the lowest (16.31 mg/L). The samples of Madison tap, spring, and Black Brook water all have similar concentrations of chloride ions ranging from 91.05 to 106.36 mg/L. Moreover, the chloride ion concentrations of the water samples taken from both ponds on 7/18 were significantly higher than the pond samples of 8/5.

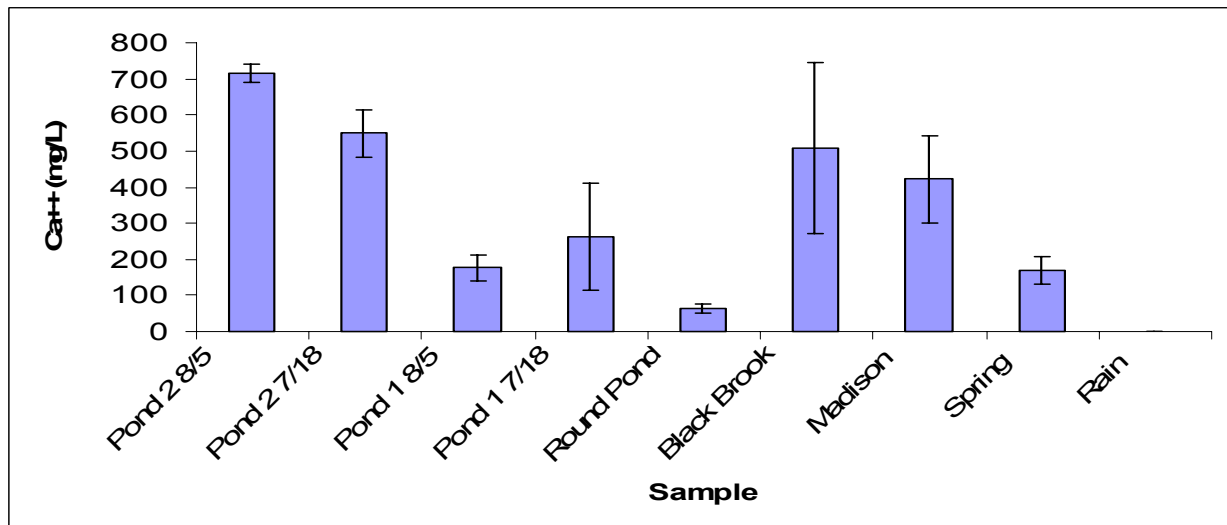Total Calcium Ion Concentration

Figure 4: Calcium ion concentration in different water samples

Figure 4 shows that the water sample taken from Pond 2 on 7/18 had the highest Ca^{2+} concentration (548.73 mg/L). Meanwhile, the Ca^{2+} concentration of rainwater was found to be negligible. In addition, the Pond 2 (7/18), Black Brook, and Madison tap water all have similar

concentrations of Ca^{2+}. Furthermore, the Ca^{2+} concentrations in both Ponds 1 and 2 decreased from 7/18 to 8/5.

DISCUSSION

Although water hardness is primarily concerned with the number of calcium ions present, other ions can also contribute to the "hardness" of water. In the calcium titrations performed, Madison tap water was found to contain 428.26 mg/L of ions, while spring water only had 167.64 mg/L of ions. Because tap water contains 260 mg/L more calcium ions than spring water, we would normally conclude that tap water is the harder water. However, in a crude titration experiment with soap water, results showed that the spring water required much more soap to titrate. Due to the function and structure of the soap molecule (see Methods), hard water requires more soap to form the same amount of bubbles. The soap molecules' sodium or potassium ion is replaced by the calcium ion, inhibiting its function to form bubbles. When all the soap molecules have bonded with the calcium ions, excess soap molecules begin to form bubbles on the surface of the solution. From this information, it can be concluded that the greater the amount of soap necessary to achieve soap bubble formation, the greater the quantity of "hard" ions present, and the harder the water. Madison water averaged 27.00 ml of soap solution to titrate, while spring water required approximately 49.33 ml. This ambiguity between the soap water test and the calcium titration can be explained by the presence of ions other than calcium in the water samples. Although the calcium content of tap water is much higher than that of spring water, the results of the soap test suggest that the content of other ions (such as magnesium and iron) are higher in the spring water than the tap water. Further evidence for this determination is the conductivity test, in which spring water had a higher ion molarity. While the soap test was the crudest test performed, there is still a significant difference between the volumes of soap solution required for tap and spring water, which cannot be dismissed by poor experimental sources of error. It is evident from these soap test results that holistically, spring water is harder than Madison tap water because of its observed greater presence of hard water ions.

Conductivity is a measure of how easily current will flow throughout a sample of water. Electricity is passed through a solution from one ion to the next, indicating that the higher the ion concentration, the stronger the electric current flow and the greater the conductivity [6]. The relationship measured between the ion molarity and the solution's conductivity is linear; it is found that the greater the ion molarity, the greater the solution's ability to conduct electricity. However, if the molarity of a particular ion increases too much in equilibrium, the rate of ionization decreases due to a saturation of the solution, and the reverse reaction and decrease in ions is favored. At this point the ions will act against each other and electric current will not be able to flow as effectively [7].
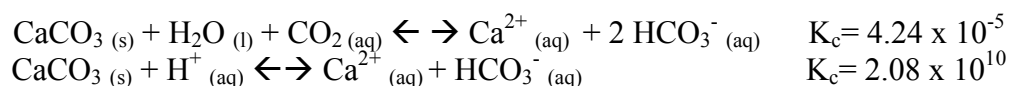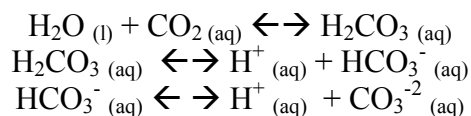
Although we determined that spring water had a higher total ion concentration than Madison tap water, Madison tap water contained higher concentrations of both the chloride and calcium ions. This suggests that there are ions other than Cl^- or Ca^{2+} that are present in larger amounts in the spring water than in the Madison water. As aforementioned, it is very probable that "hard" ions such as magnesium and iron are present in the spring water. Another possible ion is fluorine, as it is sometimes added to water to help facilitate binding of calcium to the teeth.

Additionally, spring water could contain ionic compounds of greater gram molecular weight than the compounds in tap water.

The pH readings of the various water samples not only told us the amount of hydrogen ions present, but most importantly, provided us with an indication of how readily calcium carbonate would dissociate in each type of water. In pure pH 7 water, calcium carbonate dissociates very slightly in the following net ionic equation:

$$CaCO_{3\ (s)} \rightarrow Ca^{2+}_{\ (aq)} + CO_3^{\ 2-}_{\ (aq)} \qquad K_{sp}= 4.47 \times 10^{-9}$$
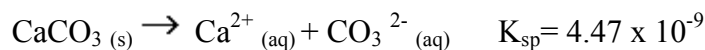
The extremely low solubility product constant, or K_{sp}, shows that calcium carbonate is only very slightly soluble in water. However, in natural environments (i.e. the water samples we tested), carbon dioxide is present in the water and must be taken into account. This presence of carbon dioxide may be due any photosynthetic processes of aquatic plant life or the double replacement combination of carbon dioxide gas with water vapor.   Water and carbon dioxide form carbonic acid, which in turn releases its hydrogen ions one at a time to ultimately form the carbonate radical. The equilibrium reactions for the dissociation of calcium carbonate then become:

$$H_2O_{\ (l)} + CO_{2\ (aq)} \longleftrightarrow H_2CO_{3\ (aq)}$$
$$H_2CO_{3\ (aq)} \longleftrightarrow H^+_{\ (aq)} + HCO_3^-_{\ (aq)}$$
$$HCO_3^-_{\ (aq)} \longleftrightarrow H^+_{\ (aq)} + CO_3^{-2}_{\ (aq)}$$

$$CaCO_{3\ (s)} + H_2O_{\ (l)} + CO_{2\ (aq)} \longleftrightarrow Ca^{2+}_{\ (aq)} + 2\ HCO_3^-_{\ (aq)} \qquad K_c= 4.24 \times 10^{-5}$$
$$CaCO_{3\ (s)} + H^+_{\ (aq)} \longleftrightarrow Ca^{2+}_{\ (aq)} + HCO_3^-_{\ (aq)} \qquad K_c= 2.08 \times 10^{10}$$

As the calcium carbonate solid reacts with the aqueous carbon dioxide and water, a small amount of the solid dissociates into calcium ions and HCO_3^- ions. However, in an acidic medium, calcium carbonate reacts very strongly with the hydrogen ion to form the calcium ion. The excess bicarbonate ions react with the excess hydrogen ions to form carbonic acid, which in turn, breaks down into water and aqueous carbon dioxide. Those products, along with calcium carbonate, become the reactants for the first reaction to occur again. It is found, therefore, that the more acidic the solution, the more readily the calcium carbonate dissociates into the calcium and carbonate ions. Nonetheless, if there is no calcium carbonate present in the various water locations, the pH does not affect the calcium ion concentration and does not contribute to the water hardness.

There is a direct correlation between the results from the pH tests and the concentration of calcium ions determined by the calcium titration. Using the pH probe, we determined that Pond 2 (8/5), Pond 1 (8/5), and Madison tap water were all substantially acidic solutions. Pond 2 had the most acidic pH of 5.94, followed by Pond 1 at 6.08 and Madison tap water with a pH of 6.11. Due to the greater presence of hydrogen ions, any calcium carbonate present in Pond 2 was dissociated more readily, subsequently becoming the water sample with the highest calcium ion concentration. Likewise, the acidic nature of Madison tap water gives rise to its high Ca^{2+} content (see Results), while the less acidic spring water yielded a less Ca^{2+} concentrated solution.

Since Pond 1 was more acidic than tap water, one would expect to see more Ca^{2+} in the pond than in the tap water. However, the titrations show that the tap water contained significantly more calcium ions per liter of water. This discrepancy could be due to a possibly higher calcium carbonate source at the Great Swamp than at the location of Pond 1. The Great Swamp is also the converging point of many local streams and ponds, which may all carry calcium ions into the Buried Valley Aquifer. Moreover, since the pH varies with temperature, the deviation could be explained by a difference in temperature between the two samples; tap water could have been slightly cooler than room temperature pond water, as it is stored in the metal water pipes.

We expected to see acidic pH readings for all of the untreated water samples taken from an outdoor source because of the possibility of acid rain. Surprisingly, the rainwater's average pH of 7.03 suggests that the air around the Great Swamp is quite clean. This pH value may be misleading because rain acidity can be affected daily by wind speed, direction, and other environmental factors. Table 4 shows us that trial 2 of the rainwater pH test measured pH 6.41, while trial 1 calculated pH 7.65. This leads us to suggest that the acidity of the brook and pond samples is due somewhat to the occasional acidity of New Jersey rain.

In general, even with increased rain and lowered pH of the water samples, the concentration of ions is comparatively less than before. This is attributed to the vast increase rain that diluted the samples. In milligrams/liter, concentration was substantially affected by the increase in the volume of new rainwater.

Through our study, we found a large amount of Ca^{2+} in the water samples, which can be accounted for by both natural and human causes. Based on the experiments performed, we concluded that no detectable amounts of calcium ions were present in rainwater. Therefore, it is possible to conclude that water does not encounter calcium until it reaches the ground.

Madison water, which ultimately comes from the mixing of various waters (including Pond 1, Pond 2, and Black Brook), had the second highest calcium ion concentration. The most probable source of the Ca^{2+} ions in the natural water samples is the dissociation of calcium carbonate.

One of the most abundant sources of calcium comes from man-made compounds. We include many different calcium-containing compounds in various types of products, and these compounds can easily enter water. For example, calcium hydroxide, $Ca(OH)_2$, is a paint additive known as "hydrated lime." It is included in many paints and coatings in order to resist the growth of odor-causing bacteria, mold and mildew. Another man-made source is road deicers [8]. Calcium chloride, $CaCl_2$, is one of the most commonly used deicers and an alternative to NaCl.

When it rains, these deicers run off into the local streams and ponds. We found that the Pond 2 (7/18), had the highest calcium ion concentration. Because it is located adjacent to an apartment complex, the great amount of snow that precipitated this past winter seems to have led to much deicing of the sidewalks, roads, and parking lots, which eventually was washed down

into the pond. When we took another sample of water from Pond2 (8/5), its calcium ion concentration had decreased due to dilution after a long month of rain.

The calcium ion concentration of Madison tap water was found to be the approximate average of Pond 2 (8/5), Pond 1 (8/5) and Black Brook. Pond 2 has a high concentration of calcium ions, but both Pond 1 and Black Brook have lower calcium levels in their water. Because all three sources feed into the Buried Valley Aquifer, it is logical that Madison tap water's calcium concentration should roughly be the arithmetic mean.

Because Round Pond has low calcium ion concentrations specifically compared to Pond 2, which is located in proximity to an apartment complex, we were able to conclude that man-made compounds and deicers are a major source of calcium ions. Very little road salt runs off into Round Pond, which would account for its lower concentration. Black Brook has greater calcium content than Round Pond because it is located near a golf course. Lime is regularly applied to maintain the grass and contributes to the amount of calcium in Black Brook.

It is not surprising that Madison tap water has the second-highest calcium concentration of all the water samples. With the majority of the local streams and ponds converging at the Great Swamp, and diffusing through the soil into the aquifer, all the road salts and other calcium sources are brought into one location. The salts, paints, etc. the town of Madison itself uses also needs to be taken into account, as Madison chemicals may also run off into these streams.

By testing water samples from five different locations, we discovered an unusually high chloride ion concentration in Pond 2 (7/18). Similarly to the calcium ion, the chloride ion levels can be explained by Loantaka Brook's proximity to apartment complexes. Contaminants may be introduced artificially from industry, failures in septic or sewer systems, pesticides and fertilizers, deicing salts, and other means.

The results obtained from chloride titrations performed on water samples from Pond 1 (7/18) determine that, like Pond 2, it also has an excessive chloride content. The 8/5 tests of both ponds confirm our predictions, as a significant decrease in the concentration of chloride and overall ion concentration in the two ponds observed due to dilution.

The observed value of chloride concentration in the Madison tap water was slightly higher than that of the spring water. Because the tap water is subjected to more extensive refining processes, most often using chlorine to kill microbiological contaminants, the chloride ion concentration increases. Alternate disinfectants may also be used, such as chlorine dioxide and chloramines.

Black Brook and Round Pond showed a moderate chloride concentration as expected due to their isolated positions in relation to other water sources. Overall, the standard deviation derived from our data collection indicated precision of results obtained from the different groups.

However, the most consistent source of error throughout the experiment was the low grade equipment used. Most of the scales used only measured to a tenth of a gram, whereas other

measured to hundredths or thousandths. The accuracy of the volumetric measurements was also compromised because the gradations of the measuring devices were large.

In general, the soap titration was not at all definitive because the soap solution had no set molarity, and there was no true equation to derive the "hardness" of the water. Bubble size and endurance were also subjective, as was the stirring size and intensity.

The water used during this experiment could also be a source of error. De-ionized water should give a reading of zero on the conductivity scale. However, it gave readings of -2 and +2. This indicates that either the water was not 100% de-ionized, or the apparatus used were not completely clean. Also, collected water varies from day to day. For example, if a large rainfall occurred the day before collection, the ions in the water would be much more diluted. Some of the water collected had large particles floating inside and was filtered before titration. However, it was not necessary to filter some water samples such as rainwater and tap water. This inconsistency may have affected the results of the titrations.

Chemicals, for the sake of time, had to be made a few days before testing. Some solutions, such as silver nitrate, change composition upon exposure to light. If this occurred, error would be expected.

Also, endpoint titration color was not standardized; therefore, resultant titrations were subjective. The endpoints varied from very light orange to a dark red hue.

CONCLUSION

Tests of water hardness, total dissolved substances, conductivity, pH, and Ca^{2+} and Cl^{-} ion concentrations enabled us to study and compare the quality of the various waters present in the Madison area. The tap water, used by every household in Madison, comes from the Buried Valley Aquifer that is located beneath the Great Swamp. This Great Swamp is fed by many water sources, including Pond 1, Pond 2, and Black Brook. Although Madison tap and spring water have very different calcium and chloride titration readings, both samples show similar conductivity, leading us to believe that the two have approximately the same amount of ions present. Due to the extent of rapid development in New Jersey, it is essential that scientists continue to study and investigate this mysterious topic.

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APPENDIX A Map of Madison and surrounding area. Brooks of study are highlighted and sample collection sites are marked as dots.