Determining Water Hardness
by EDTA Titration

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Purpose of the Experiment
Determine the total hardness, or total calcium and magnesium ion concentration, in a water sample by titrating with EDTA solution. After appropriate treatment of the water, determine the permanent hardness, the temporary hardness, and the individual concentrations of calcium ion and magnesium ion.

Background Information

Natural Water and Water Hardness
Natural water dissolves many organic salts as it flows through rocks and soil. Water containing significant amounts of calcium ion (Ca$^{2+}$), magnesium ion (Mg$^{2+}$), iron(II) ion (Fe$^{2+}$ or ferrous ion), or iron(III) ion (Fe$^{3+}$ or ferric ion) is known as hard water. Ca$^{2+}$ and Mg$^{2+}$ ions are the most common sources of hardness in water.

Hardness in water causes problems at home as well as in industry. Ca$^{2+}$ and Mg$^{2+}$ ions react with soap to form water-insoluble salts. As a result, the soap loses some of its cleaning power, and the insoluble Ca$^{2+}$ and Mg$^{2+}$ salts of fatty acids form a scum that sticks to sinks, bathtubs, and fabrics. When hard water containing relatively large amounts of hydrogen carbonate ion (HCO$_3^-$) is heated, insoluble calcium, magnesium, and iron(II) carbonates (CaCO$_3$, MgCO$_3$, FeCO$_3$) precipitate as “boiler scale” inside pipes and vessels such as hot water heaters, teakettles, and commercial boilers.

Chemists have developed effective methods for removing the undesirable ions using a process known as water softening. For example, detergents remain soluble in the presence of Ca$^{2+}$ and Mg$^{2+}$ ions, so they are an excellent replacement for soap when washing in hard water is unavoidable. Ion exchangers such as zeolites substitute sodium ions (Na$^+$) for Ca$^{2+}$ and Mg$^{2+}$ ions in hard water. Water that has passed through such an ion exchanger is often called soft water. Synthetic ion-exchange resins purify or “demineralize” hard water by exchanging both positive and negative ions.

We use data obtained from water hardness analyses to aid our selection of appropriate methods for reducing the hardness of water from a specific source. The total hardness of the water refers to the total concentration of Ca$^{2+}$ and Mg$^{2+}$ ions in the water. The temporary hardness, or carbonate hardness, of the water refers to the amount of Ca$^{2+}$ and Mg$^{2+}$ ions that can be removed as insoluble carbonates by boiling the water. Equations 1 and 2 show the formation of these insoluble carbonates, which precipitate when carbon(IV) oxide gas (CO$_2$, or carbon dioxide) is driven off by heat.
Ca\(^{2+}\) (aq) + 2 HCO\(_3^-\) (aq) \(\xrightarrow{\text{heat}}\) CaCO\(_3\) (s) + H\(_2\)O(l) + CO\(_2\) (g) \hspace{1cm} \text{(Eq. 1)}

Mg\(^{2+}\) (aq) + 2 HCO\(_3^-\) (aq) \(\xrightarrow{\text{heat}}\) MgCO\(_3\) (s) + H\(_2\)O(l) + CO\(_2\) (g) \hspace{1cm} \text{(Eq. 2)}

We determine the amount of temporary hardness in water by the HCO\(_3^-\) ion concentration in the water and the stoichiometry of Equations 1 and 2. Whenever the concentration of HCO\(_3^-\) ion is less than twice the total concentration of the Ca\(^{2+}\) and Mg\(^{2+}\) ions, precipitation of these cations is incomplete. In this experiment, permanent hardness, or non-carbonate hardness, refers to the amount of Ca\(^{2+}\) and Mg\(^{2+}\) ions remaining in solution after removal of temporary hardness. Total hardness is the sum of permanent and temporary hardness.

**EDTA Titration Methods**

We can determine the total amount of Ca\(^{2+}\) and Mg\(^{2+}\) ions in water by titrating a water sample with a standardized ethylenediaminetetraacetic acid (EDTA) solution, using an organic dye indicator called Eriochrome Black T (EriT). EDTA is a weak acid that has the structure shown in Figure 1(a).

![Figure 1](image1.png)

**Figure 1** (a) The structure of ethylenediaminetetraacetic acid, EDTA (b) The structure of the tetracarboxylate ion, [EDTA]\(^{4-}\), formed by dissociation of EDTA

In an aqueous solution buffered at pH 10, EDTA dissociates to form a maximum amount of tetracarboxylate ion, [EDTA]\(^{4-}\), that has the structure shown in Figure 1(b). This ion is electron-rich, having six bonding sites: the four carboxylate groups and the two nitrogen atoms. Each site has an electron pair available for bonding. The [EDTA]\(^{4-}\) anion wraps itself around a Ca\(^{2+}\) or Mg\(^{2+}\) ion so that all six electron pairs are shared with the metal ion, as shown in Figure 2. In this manner, [EDTA]\(^{4-}\) forms strong 1:1 complexes, known as chelates, with metal ions such as Ca\(^{2+}\) and Mg\(^{2+}\). The word chelate derives from the Greek word for “claw.”

![Figure 2](image2.png)

**Figure 2** Structure of [Mg–EDTA]\(^{2-}\) chelate

In an aqueous solution buffered at pH 10, EriT also dissociates, forming [H–EriT]\(^{2-}\) ion, a blue ion that bonds with either Mg\(^{2+}\) or Ca\(^{2+}\) ion to form a wine-red complex. The reactions of [H–EriT]\(^{2-}\) ion with Ca\(^{2+}\) and Mg\(^{2+}\) ions are reversible and are shown in Equations 3 and 4.

\[
\text{Ca}^{2+} (\text{aq}) + [\text{H–EriT}]^{2-} (\text{aq, blue}) + \text{H}_2\text{O}(l) \xrightarrow{\text{pH 10}} [\text{Ca–EriT}]^{+} (\text{aq, wine-red}) + \text{H}_3\text{O}^+ (\text{aq}) \hspace{1cm} \text{(Eq. 3)}
\]

\[
\text{Mg}^{2+} (\text{aq}) + [\text{H–EriT}]^{2-} (\text{aq, blue}) + \text{H}_2\text{O}(l) \xrightarrow{\text{pH 10}} [\text{Mg–EriT}]^{+} (\text{aq, wine-red}) + \text{H}_3\text{O}^+ (\text{aq}) \hspace{1cm} \text{(Eq. 4)}
\]

Complexes of Ca\(^{2+}\) and Mg\(^{2+}\) ions with [EDTA]\(^{4-}\) and [H–EriT]\(^{2-}\) ions have different stabilities. The order of decreasing stability of these complexes is:

\[
[\text{Ca–EDTA}]^{2-} > [\text{Mg–EDTA}]^{2-} > [\text{Mg–EriT}]^{-} > [\text{Ca–EriT}]^{-}
\]

If we add several drops of EriT indicator solution to a properly buffered aqueous solution containing both Ca\(^{2+}\) and Mg\(^{2+}\) ions, the small amount of [Mg–EriT]\(^-\) complex which forms turns the solution wine-red. The wine-red color persists as we add EDTA solution, until we reach the titration end point. As we titrate the Ca\(^{2+}\) and Mg\(^{2+}\) ions with EDTA solution, the titration reactions shown in Equations 5–7 occur in order:
Ca\(^{2+}\)(aq) + [EDTA]\(^{4-}\)(aq) \rightarrow [Ca–EDTA]\(^{2-}\)(aq)  
(Eq. 5)

Mg\(^{2+}\)(aq) + [EDTA]\(^{4-}\)(aq) \rightarrow [Mg–EDTA]\(^{2-}\)(aq)  
(Eq. 6)

[Mg–ErioT]\(^{-}\)(aq, wine-red) + [EDTA]\(^{4-}\)(aq) + H\(_3\)O\(^{+}\)(aq) \rightarrow [Mg–EDTA]\(^{2-}\)(aq) + [H–ErioT]\(^{2-}\)(aq, blue) + H\(_2\)O(l)  
(Eq. 7)

We reach the titration end point when the reactions in Equations 5 and 6 are complete and the [EDTA]\(^{4-}\) anion finally breaks up the wine-red [Mg–ErioT]\(^{-}\) complex, releasing blue [H–ErioT]\(^{2-}\) ion. The solution changes from wine-red, through a transitory purple, to a distinct, permanent blue.

Conventions for Reporting Water Hardness

We commonly report water hardness as milligrams of equivalent CaCO\(_3\) per liter of water or as parts of equivalent CaCO\(_3\) per million parts of water (ppm of equivalent CaCO\(_3\)). In other words, we determine the number of millimoles of Ca\(^{2+}\) and Mg\(^{2+}\) ions present in one liter of water, and then assume that all of these cations originate from dissolved CaCO\(_3\). We calculate the mass of this hypothetical CaCO\(_3\) and use this mass when reporting our results.

Recall that 1 g = 1000 mg, 1 L = 1000 mL, and 1 mol = 1000 mmol. We use our titration data and Equations 8 and 9 to calculate the total number of millimoles of Ca\(^{2+}\) and Mg\(^{2+}\) ion present in our carefully measured water sample.

\[
\text{number of millimoles of Ca}^{2+} \text{ and Mg}^{2+} \text{ ions, mmol} = \frac{\text{number of millimoles of Ca}^{2+} \text{ and Mg}^{2+} \text{ ions, mmol}}{\text{volume of EDTA solution required, L}} \times \frac{\text{concentration of EDTA solution, mol L}^{-1}}{\text{mol L}^{-1}} 
\]  
(Eq. 8)

\[
\text{number of millimoles of Ca}^{2+} \text{ and Mg}^{2+} \text{ ions, mmol} = \frac{\text{number of millimoles of Ca}^{2+} \text{ and Mg}^{2+} \text{ ions, mmol}}{\text{number of moles of EDTA required, mol}} \times \frac{1000 \text{ mmol}}{1 \text{ mol}} 
\]  
(Eq. 9)

Suppose we find that 100 mL of water contains a total of 0.150 mmol of Ca\(^{2+}\) and Mg\(^{2+}\) ions. We use these data and Equation 10 to calculate the total metal ion concentration in units of millimoles per liter of water.

\[
\text{total Ca}^{2+} \text{ and Mg}^{2+} \text{ ion concentration, mmol (L H}_2\text{O)}^{-1} = \frac{\text{number of millimoles of Ca}^{2+} \text{ and Mg}^{2+} \text{ ions in sample, mmol}}{\text{water sample volume, mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} 
\]  
(Eq. 10)

\[
= \frac{0.150 \text{ mmol}}{100 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 1.50 \text{ mmol (L H}_2\text{O)}^{-1} 
\]

Next, we assume that the total concentration of Ca\(^{2+}\) and Mg\(^{2+}\) ions is derived from dissolved CaCO\(_3\). In other words, we assume that the total hardness of 1.50 mmol of Ca\(^{2+}\) and Mg\(^{2+}\) ions per liter of water is equivalent to 1.50 mmol of dissolved CaCO\(_3\) per liter of water. We then use the molar mass of CaCO\(_3\) (100 g mol\(^{-1}\)) and Equation 11 to convert millimoles of equivalent CaCO\(_3\) per liter of water to milligrams of equivalent CaCO\(_3\) per liter of water.

\[
\text{number of milligrams of equivalent CaCO}_3 = \frac{\text{number of millimoles of equivalent CaCO}_3 \text{ per liter, mmol equiv. CaCO}_3 (\text{L H}_2\text{O})^{-1} \times \frac{1 \text{ mol CaCO}_3}{1000 \text{ mmol}}}{\text{mmol equiv. CaCO}_3 (\text{L H}_2\text{O})^{-1} \times \frac{100 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3}} 
\]

\[
= \frac{1.50 \text{ mmol of equiv. CaCO}_3 (1 \text{ mol})}{1000 \text{ mmol}} \times \frac{100 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3} \times \frac{1000 \text{ mg}}{1 \text{ g}} 
\]

\[
= 1.50 \times 10^2 \text{ mg equiv. CaCO}_3 (\text{L H}_2\text{O})^{-1} 
\]

Because the density of water is 1.00 g mL\(^{-1}\), we can conveniently express the number of milligrams of equivalent CaCO\(_3\) per liter of water as the number of milligrams of equivalent CaCO\(_3\) per 1.00 \times g of water, or as the number of parts of equivalent CaCO\(_3\) per million parts of water (ppm of equivalent CaCO\(_3\)), as shown in Equation 12.

\[
\text{number of milligrams of equiv. CaCO}_3 \text{ per 1.0 } \times 10^6 \text{ mg water, or ppm equiv. CaCO}_3 = \frac{\text{number of milligrams of equiv. CaCO}_3, mg}{1 \text{ L H}_2\text{O}} \times \frac{1 \text{ L H}_2\text{O}}{1000 \text{ mL H}_2\text{O}} \times \frac{1.00 \text{ mL H}_2\text{O}}{1000 \text{ mg}} 
\]  
(Eq. 12)
\[
= \left( \frac{1.50 \times 10^2 \text{ mg}}{1 \text{ L} \text{ H}_2\text{O}} \right) \left( \frac{1 \text{ L H}_2\text{O}}{1000 \text{ mL H}_2\text{O}} \right) \times \left( \frac{1.00 \text{ mL H}_2\text{O}}{1 \text{ g} \text{ H}_2\text{O}} \right) \left( \frac{1 \text{ g}}{1000 \text{ mg}} \right)
\]

\[= 1.50 \times 10^2 \text{ ppm of equivalent CaCO}_3\]

We can use similar calculations to express individual Ca\(^{2+}\) and Mg\(^{2+}\) ion concentrations in terms of ppm of equivalent CaCO\(_3\). For instance, we report 0.100 mmol of Ca\(^{2+}\) ion in 100 mL of water (1.00 mmol L\(^{-1}\)) as 100 ppm of equivalent CaCO\(_3\). We report 5.0 \(\times\) 10\(^{-2}\) mmol of Mg\(^{2+}\) ion in 100 mL of water (0.50 mmol L\(^{-1}\)) as 50 ppm of equivalent CaCO\(_3\).

Finally, we determine the individual Ca\(^{2+}\) and Mg\(^{2+}\) ion concentration in the water sample. We do this by quantitatively precipitating Ca\(^{2+}\) ion as calcium oxalate (CaC\(_2\)O\(_4\)) from a known volume of the water, using Ca\(^{2+}\) ion-precipitating buffer solution. The buffer solution contains ammonium oxalate, (NH\(_4\))\(_2\)C\(_2\)O\(_4\), ammonium chloride (NH\(_4\)Cl) and ammonium hydroxide (NH\(_4\)OH). The formation of CaC\(_2\)O\(_4\) is shown in Equation 13.

\[
\text{Ca}^{2+}(\text{aq}) + \text{C}_2\text{O}_4^{2-}(\text{aq}) \rightarrow \text{CaC}_2\text{O}_4(\text{s, white})
\] (Eq. 13)

Because CaC\(_2\)O\(_4\) is very insoluble in water, the addition of excess oxalate ion (C\(_2\)O\(_4^{2-}\)) will precipitate the maximum amount of Ca\(^{2+}\) ion possible so that it will not interfere with the next titration. We remove the precipitate by filtration, then analyze the filtrate for Mg\(^{2+}\) ion by titrating with EDTA solution. The difference between calculated total hardness and Mg\(^{2+}\) ion hardness is the Ca\(^{2+}\) ion hardness in the water sample.

Summary of Experimental Determinations and Calculations

In this experiment, you will determine the total hardness of a water sample by titrating a known volume of the water with a standardized EDTA solution, using Eriot indicator. From the molarity and volume of EDTA solution used and the volume of water titrated, you will calculate the total Ca\(^{2+}\) and Mg\(^{2+}\) ion concentration in your water sample. You will report this total hardness in ppm of equivalent CaCO\(_3\).

You will also determine the permanent hardness and temporary hardness of your water sample. To do so, you will boil a known volume of the water, filter any precipitated CaCO\(_3\) and MgCO\(_3\), and analyze the filtrate for permanent hardness by titrating with EDTA solution. You will use your experimental results from Parts III and IV to calculate the temporary hardness in your water sample.

Finally, you will determine the individual Ca\(^{2+}\) and Mg\(^{2+}\) ion concentration in your water sample, using your titration data.

Procedure

Chemical Alert

pH 10 buffer solution (contains ammonium chloride and ammonium hydroxide)—toxic, corrosive, and irritant

Ca\(^{2+}\)ion–precipitating buffer solution (contains ammonium chloride, ammonium hydroxide, and ammonium oxalate)—toxic, corrosive, and irritant

Caution: Wear departmentally approved eye protection while doing this experiment.

I. Preparing a Water Sample for Mg\(^{2+}\) Ion Analysis

Note: Your laboratory instructor will demonstrate a satisfactory method for measuring and transferring your water samples.

Note: The fumes from the NH\(_4\)OH in the Ca\(^{2+}\)ion–precipitating buffer are toxic and irritating. Dispense the buffer solution in a well-ventilated place or under a fume hood.

Transfer 75 mL of your water sample into a clean, dry 250-mL beaker. Add 20 mL of Ca\(^{2+}\) ion-precipitating buffer solution, and mix well. Cover the beaker with a watch glass, label it “Beaker 1,” and set the mixture aside to stand for at least 1 hr. Proceed to Part II.

Note: Your laboratory instructor may suggest that you boil the water sample described in Part IV while you are preparing your buret for the titrations. Instructions for doing so appear in Part IVA, Preparing the Sample.
II. Preparing the Buret for Titration

Rinse and fill a 50-mL buret with standardized EDTA solution according to the directions of your laboratory instructor. Drain a small amount of the EDTA solution from the buret in order to remove any air bubbles and to fill the tip. Record the exact EDTA concentration and your initial buret reading, to the nearest 0.02 mL, on Data Sheet 1.

III. Determining the Total Hardness of the Water Sample

Note: The fumes from the NH₄OH in the pH 10 buffer solution are toxic and irritating. Dispense the buffer solution in a well-ventilated place or under a fume hood.

Obtain 75 mL of your water sample in a second clean, dry 250-mL beaker. Transfer 25.00 mL of the water into a clean, 125-mL Erlenmeyer flask. Using droppers, add 1 mL of pH 10 buffer solution and 2 drops of 1% Eriostain indicator solution to the flask. Gently swirl the contents of the flask until the solution is uniformly wine-red throughout. Without delay, titrate the water with EDTA solution until the color changes from wine-red to permanent blue. Near the end point, add EDTA solution dropwise, swirling the flask and its contents for 5 to 10 s after each addition of EDTA solution. The indicator color change requires a few seconds to develop. Record your final buret reading on Data Sheet 1.

Repeat this procedure two more times, each time transferring a fresh 25.00-mL sample from Flask IV into a clean 125-mL Erlenmeyer flask. Pour any solution remaining in the volumetric flask into the drain. Pour the titration mixtures into the “Discarded Titration Mixtures” container. Rinse the Erlenmeyer and volumetric flasks with distilled water and drain to dry.

IV. Determining the Mg²⁺ Ion Hardness

Prepare and fill the buret with standardized EDTA solution as described in Part II. Record your initial buret reading on Data Sheet 3.

A. Preparing the Sample

Using a filtering funnel fitted with filter paper, filter the mixture prepared in Part I directly into a clean, dry 100-mL volumetric flask. Rinse the beaker and precipitate with several 1-mL portions of distilled water. Add the rinses to the volumetric flask by pouring them through the funnel. Rinse all Erlenmeyer flasks with distilled water and drain to dry.

B. Titrating the Sample

Transfer 25.00 mL of the solution from Flask V into a clean, dry 125-mL Erlenmeyer flask. Add pH 10 buffer solution and 1% Eriostain indicator solution to the sample in the Erlenmeyer flask, as described in Part III. Titrate the resulting wine-red solution with your standardized EDTA solution also as described in Part III. Record your final buret reading on Data Sheet 3.
Repeat the titration two more times, each time transferring a fresh 25.00-mL sample from Flask V into a clean 125-mL Erlenmeyer flask.

Pour the titration mixtures, any EDTA solution remaining in your buret, and any solution remaining in the volumetric flasks into the “Discarded Titration Mixtures” container.

**Caution:** Wash your hands thoroughly with soap or detergent before leaving the laboratory.

### Calculations

Do the following calculations for each determination and record the results on the appropriate sections of your Data Sheets.

1. For each analysis in Parts III, IV, and V, calculate the number of moles and millimoles of EDTA required for titration, using Equations 8 and 9.

2. For each analysis, calculate the mean number of millimoles of EDTA required.

3. For the determination of the total hardness of your water sample, Part III, calculate:
   - (1) the total hardness of the water in the total number of millimoles Ca\(^{2+}\) and Mg\(^{2+}\) ions per liter of water. Use Equation 10 and the mean number of millimoles of EDTA required, noting that the number of millimoles of EDTA is equal to the total number of millimoles of Ca\(^{2+}\) and Mg\(^{2+}\) ions in the titrated water sample.
   - (2) the total hardness of the water in ppm of equivalent CaCO\(_3\). Use Equations 11 and 12.

4. For the determination of the permanent and temporary hardness of your water sample, Part IV, calculate:
   - (1) the permanent hardness in millimoles of Ca\(^{2+}\) and Mg\(^{2+}\) ions per liter of water. Use Equation 10 and the mean number of millimoles of EDTA required, noting that permanent hardness is the number of millimoles of Ca\(^{2+}\) and Mg\(^{2+}\) ions that remain in solution after the precipitated CaCO\(_3\) has been removed.
   - (2) the permanent hardness of the water in ppm of equivalent CaCO\(_3\). Use Equations 11 and 12.
   - (3) the temporary hardness of the water in ppm of equivalent CaCO\(_3\), using your experimental results from Parts III and IV.

5. From the separate Mg\(^{2+}\) ion analysis (Part V), calculate:
   - (1) the mean number of millimoles of Mg\(^{2+}\) ions per liter of filtrate. Use the mean number of millimoles of EDTA required in Part V, noting that Ca\(^{2+}\) ions were quantitatively removed from the sample prior to titration.
   - (2) the mean number of millimoles of Mg\(^{2+}\) ions per liter of your water sample. Remember to correct for the dilution from 75 mL to 100 mL.
   - (3) the hardness due to Mg\(^{2+}\) ions in your water sample, in ppm of equivalent CaCO\(_3\).
   - (4) the hardness due to Ca\(^{2+}\) ions in the water, in ppm of equivalent CaCO\(_3\), noting that the total hardness is equal to the hardness due to Mg\(^{2+}\) ions plus the hardness due to Ca\(^{2+}\) ions.
Post-Laboratory Questions

(Use the spaces provided for the answers and additional paper if necessary.)

1. You adjusted the filtrate volume in Part IV of this experiment with distilled water prior to taking a sample for titration with EDTA solution. Briefly explain:
   (1) why the dilution to a known volume was necessary.
   (2) why the addition of distilled water did not cause an error in your determination of the permanent hardness of the water.

2. Calcium carbonate frequently deposits in home water pipes. Should a homeowner expect more CaCO₃ buildup in hot water pipes or cold water pipes? Briefly explain.

3. EDTA and other similar chelating agents are frequently used in cleaning products.
   (1) Briefly explain how a cleaning product containing EDTA can aid in removing hard water soap scum from a sink.
   (2) Would the presence of EDTA in a soap powder help prevent soap scum formation in hard water? Briefly explain.
Data Sheet 1

III. Determining the Total Hardness of the Water Sample

source of water sample

volume of water sample titrated, mL

concentration of EDTA solution, mol L⁻¹

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<td>total hardness of water sample, ppm of equivalent CaCO₃</td>
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Data Sheet 2

IV. Determining the Permanent and Temporary Hardness of the Water Sample

volume of water boiled, mL

adjusted volume of filtrate, mL

volume of filtrate titrated, mL

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Data Sheet 3

V. Determining the Mg$^{2+}$ Ion Hardness

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<th>determination 3</th>
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<td>number of millimoles of EDTA required, mmol</td>
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|                       |                  |                  |                  |
| mean number of millimoles of EDTA required, mmol |                  |                  |                  |
| mean number of millimoles of Mg$^{2+}$ ions per liter of filtrate, mmol |                  |                  |                  |
| mean number of millimoles of Mg$^{2+}$ ions per liter of water sample, mmol |                  |                  |                  |
| hardness due to Mg$^{2+}$ ions in water sample, ppm of equivalent CaCO$_3$ |                  |                  |                  |
| total hardness of the water (from Part III), ppm of equivalent CaCO$_3$ |                  |                  |                  |
| hardness due to Ca$^{2+}$ ions in water sample, ppm of equivalent CaCO$_3$ |                  |                  |                  |
Pre-Laboratory Assignment

1. Briefly describe the hazards associated with the Ca\textsuperscript{2+} ion-precipitating buffer solution you will use in this experiment (in Part I).

2. Briefly explain what is meant by the statement, “Ca\textsuperscript{2+} ion is quantitatively removed from solution by precipitation with C\textsubscript{2}O\textsubscript{4}\textsuperscript{2–} ion.”

3. Briefly distinguish between the terms temporary water hardness and permanent water hardness as they pertain to this experiment.

4. Suppose that a liter of sample water contains 1.00 mmol of Ca\textsuperscript{2+} ions and 1.50 mmol of HCO\textsubscript{3}– ions.
   (1) How many millimoles of CaCO\textsubscript{3} will precipitate if this water is boiled for 15 min?
   (2) How many millimoles of Ca\textsuperscript{2+} ions will remain in solution after the water has been boiled?

5. A student collected the following data while using the procedure described in this module: 27.13 mL of 1.00 × 10\textsuperscript{–2}M EDTA solution was required to reach the ErioT end point for titration of 25.00 mL of the sample water.

   After boiling and filtering 100.0 mL of the water, the student diluted the filtrate to 100.0 mL with distilled water. To titrate 25.00 mL of the diluted filtrate, 26.40 mL of 1.00 × 10\textsuperscript{–2}M EDTA solution was required.

   The student precipitated the Ca\textsuperscript{2+} ion as CaC\textsubscript{2}O\textsubscript{4} from 75.0 mL of the water sample and removed the precipitate by filtration. The student then adjusted the water sample volume to 100.0 mL. To titrate 25.00 mL of the diluted filtrate, 7.43 mL of 1.00 × 10\textsuperscript{–2}M EDTA solution was required.

   (1) Calculate the total number of millimoles of Ca\textsuperscript{2+} and Mg\textsuperscript{2+} ions present in the titrated sample.

   (2) Calculate the total water hardness, expressed in ppm of equivalent CaCO\textsubscript{3}.

   (3) Calculate the permanent water hardness in terms of total mmol of Ca\textsuperscript{2+} and Mg\textsuperscript{2+} ions.
(4) Express the permanent water hardness in terms of ppm of equivalent CaCO₃.

(5) Calculate the temporary water hardness in terms of ppm of equivalent CaCO₃.

(6) Calculate the number of millimoles of Mg²⁺ ions per liter of the water.

(7) Express the water hardness due to Mg²⁺ ions in terms of ppm of equivalent CaCO₃.

(8) Express the water hardness due to Ca²⁺ ions in terms of ppm of equivalent CaCO₃.