AP Chemistry Lab 07

Investigation 3 *What Makes Hard Water Hard?*

**Central Challenge**

Samples of tap water have been collected from faculty members who live in different townships, villages, cities, and counties. Some are from municipal water sources, others are from wells. The water will be analyzed for their quantities of water hardness through principles of metal ion precipitation and separation. The samples will then be ranked in order of increasing water hardness.

**Context for This Investigation**

When relocating, there are many factors that go into choosing where to buy a house. One of those factors is water quality – it is advantageous to find a location where a water softener will *not* be needed. Hard water costs more for detergent, plumbing repairs, and appliance maintenance and replacement. Having water that is already soft is preferred to installing a water softening system. You, the students, will identify which of the areas around Western New York have the lowest hard water quantities by ranking the hardness of the water in the areas.

**Materials**

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| --- | --- | --- |
| Distilled water | Metal scoops | 250 mL Erlenmeyer flasks (optional) |
| Calcium chloride, anhydrous (for practice) | 250 mL beakers | Buchner funnels with stoppers |
| Sodium carbonate, anhydrous (for practice) | 100 mL graduated cylinders | Glass tubing (optional) |
| Wash bottles | Watch glasses | Ring stand |
| Analytical balances | 250 mL side-arm flasks (filter flasks) | Drying oven |
| Clamps | Aspirators | Filter papers |
| Double hole rubber stoppers (optional) |  | Water samples |
|  |  |  |

**Safety and Disposal**

All solutions can be flushed down the drain. Normal laboratory precautions, including wearing goggles at all times, should be taken as these solutions may be harmful if swallowed and can irritate the eyes.

**Prelab Guiding Questions/Simulations**

The following seven questions are homework. Students share their responses as a lab group. Additional tips for discussion are included in the student answers.

**1.** Questions a–g relate to the interactive simulation, a PhET simulation designed by the University of Colorado. Go to *http://phet.colorado.edu/en/simulation/soluble-salts* to open the interactive simulation. Click *Run Now*. After obtaining access to the simulation, note the three tabs at the top: *Table Salt*, *Slightly Soluble Salts*, and *Design a Salt*.

1. Under the *Table Salt* tab, shake the salt shaker. Describe what happens to the solid table salt, NaCl.
2. Click *Reset All*. Shake the salt shaker until some of the particles are designated as *Bound*. How many sodium ions are designated as *Dissolved*? How many sodium ions are designated as *Bound*? Use the simulation to describe what *bound* means.
3. Click the *Slightly Soluble Salts* tab. Using the pull-down menu, select *Mercury(II) Bromide*. Slowly shake the salt shaker until some of the ions are designated as *Bound.* How many shakes did it take? Compare how this mercury(II) bromide is different from table salt.
4. Shake a large amount of mercury(II) bromide into the container. How do the number of dissolved ions change as more mercury(II) bromide is added to the container?
5. Slowly drain some of the mixture out of the container and stop. Where do the dissolved ions go as the solution is drained? What else do you notice as the mixture is drained? If the mixture left the container through a long pipe as it was drained, how might problems arise inside the pipe?
6. How might the simulation look different if the mercury(II) bromide was created from two salts, such as mercury(II) nitrate and sodium bromide, rather than added directly?
7. Predict an appropriate experimental method to collect the bound mercury(II) bromide.

**2.** Watch the following video:

*http://www.youtube.com/watch?v=YcZSNcaHHN8&feature=youtube\_gdata\_player*

After watching the video, describe what you believe to be “hard water.”

**~~3.~~** ~~Watch the following animation (hit the ‘next’ button at the bottom of the page to get to the animation):~~

*~~http://bcs.whfreeman.com/chemcom5e/content/cat\_010/Unit1\_Media/CC\_5e\_U1\_SecD.swf~~*

~~After watching this animation, explain how soap scum forms. What ions contribute~~

~~to the formation of soap scum?~~

**4.** Describe some ways that water can be softened. Some helpful sites include, but are not limited to:

*http://www.chem1.com/CQ/hardwater.html*

*http://www.ag.ndsu.edu/pubs/h2oqual/watsys/ae1031w.htm*

**5.** A reaction occurs between solutions of strontium bromide and silver nitrate, as shown in the equation below:

SrBr2 (*aq*) + AgNO3 (*aq*) → Sr(NO3)2 (*aq*) + AgBr (*s*)

**a.** If 3.491 grams of the precipitate is formed, how many moles of strontium bromide were reacted?

**b.** If 45.61 mL of strontium bromide were reacted in Part a, what is the molarity of the strontium bromide solution that was used?

**c.** In collecting the precipitate, why would it be inappropriate to heat the reacted mixture and evaporate off the water?

**6.** Below is a table of solubility product constant (*Ksp*) values. Consider how the values in this table may help in deciding how to remove one of these ions by selective precipitation. Decide which lab group member is assigned one of the following cations: **Mg2+**, **Ca2+**, **Fe2+**. How would you remove the ion from hard water? Include which anion you would use to remove the cation and explain why you chose that anion.



**7.** Using the SDS (formerly MSDS), identify the potential health risks and the appropriate measures for first aid for the following chemicals:

1. calcium chloride

**b.** sodium chloride

**c.** sodium carbonate, anhydrous

**Explanation to Strengthen Student Understanding**

Have you ever had difficulty lathering soap or find that the scum in your shower constantly needs to be removed? These are signs of “hard water.” Soap does not lather well in hard water because metal ions, such as Ca2+, form precipitates, creating “soap scum.” A *precipitate* is an insoluble compound that forms when soluble ions in separate solutions are mixed together. Because this happens, soap is a less effective cleanser in hard water. Even laundry can appear dingy or feel rough when washed in hard water.

While these metal ions are generally harmless, hard water has other disadvantages, such as “boiler scale.” Boiler scale is a scaly buildup of calcium carbonate, CaCO3, produced when the calcium ions in hard water have precipitated with dissolved carbonate ions, CO32–. This scale can build up on the inside of water pipes and coffee makers. One of the biggest problems boiler scale creates is that it reduces the operating efficiency in water heaters. Even a thin layer of scale inside a water heater can reduce the energy efficiency by 10 percent or more. Scale can also result in the failure of boiler tubes as they become clogged. Once these insoluble salts form a deposit, other metal ions present in the water can become bound to the deposit, increasing the thickness of the boiler scale layer.

The best way to control the formation of boiler scale is through water pretreatment, such as installing a water softener. Water softeners typically replace the Ca2+ ions with soluble Na+ salts. Other water softeners cause the calcium carbonate to form before the water is circulated to the water heater. Because metal ions enter water when it travels through rocks and soil en route to a home, the amount of water hardness varies. When the hard water within your home is 120–150 mg/L as CaCO3, it is recommended that a water softener is installed. If the hardness falls between 60–120 mg/L as CaCO3, the hardness is considered acceptable (moderately hard) and no water softener is needed.

Hard water can contain various metal ions, including Ca2+, Mg2+, and Fe3+. In order to fully determine the hardness of the water, each ion must be isolated separately. When a particular ion species, an *analyte*, needs to be isolated, it is possible to use the tendency of that ion to form an insoluble compound by a precipitation reaction.

When the analyte ion is formed into a precipitate, it can be collected through a process called *gravimetric analysis*, during which the precipitate is isolated, purified, dried, and weighed. From the mass and the known composition of the precipitate, the amount of the analyte in the original solution can be calculated stoichiometrically. When done properly, gravimetric analysis provides an extremely precise quantitative analysis of the analyte. Since hard water is commonly expressed as the milligrams of CaCO3 per liter of solution, the quantity of Ca2+ in the water identifies how hard the water is. This analyte can be isolated by mixing it with a solution of Na2CO3 to form the slightly soluble salt CaCO3.

Ca2+(*aq*) + Na2CO3(*aq*) → 2 Na+(*aq*) + CaCO3(*s*)

When completing a gravimetric analysis, an important consideration is that the analyte is completely precipitated. This can be accomplished by ensuring that the analyte acts as the limiting reactant in the precipitation reaction. Once the salt has precipitated, it can be collected through filtration. All of the impurities should be removed from the precipitate through washing and drying.

**Practice with Instrumentation and Procedure**

Before beginning the lab, students should complete the procedure described below. Questions pertaining to the procedure follow the activity. Students should complete the procedure in their lab groups. The questions should then be completed individually, as a homework assignment. The lab group, as a whole, should discuss the questions to help facilitate student knowledge before designing procedures together in lab groups.

**Prelab Instrumentation Procedure**

In this part of the lab, the students will make a solution of sodium carbonate and mix it with a solution of calcium chloride. The students will need to read the following procedure and create a data table for the data they will need to collect.

**1.** Weigh about 2 g of sodium carbonate in a clean, dry beaker.

**2.** Weigh about 2 g of calcium chloride in a second clean, dry beaker.

**3.** To each beaker, add about 20 mL of distilled water. Stir each until each solid is dissolved.

**4.** Pour a small amount of the sodium carbonate solution into the beaker containing the calcium chloride solution. Stir and observe the mixture. Add some more of the sodium carbonate solution while stirring and observing. After all of the sodium carbonate solution has been added, continue to stir this mixture for another couple of minutes.

**5.** Weigh a piece of filter paper.

**6.** Set up the filtering apparatus.

**7.** Insert the filter paper into the filtering apparatus. Wet the filter paper with distilled water.

**8.** Pour the contents of your beaker slowly into the funnel. Be careful as you pour so that none of the mixture flows out of the filter paper or the funnel. Use a wash bottle of water to rinse the precipitate out of the beaker with small quantities of water. Use a little more deionized water to wash the precipitate that is now collected in the filter paper.

**9.** Using a grease pencil or overhead marker, label a clean, dry watch glass with your name. Weigh the watch glass.

**10.** Carefully remove the filter paper with the precipitate from the filtering apparatus and set it on the watch glass. The filtrate can be poured down the drain.

**11.** Place the watch glass with the filter paper on it into the drying oven, which should be set between 110°C and 120°C.

**12.** Allow the filter paper to dry for 10–15 minutes. Answer Questions 1–2 on the next page while you wait.

**13.** Carefully remove the warm watch glass. Use a metal scoop to break the precipitate into small pieces.

**14.** Return the watch glass to the drying oven for another 5 minutes.

**15.** Carefully remove the watch glass out of the oven and set aside to cool. Once cool, weigh the watch glass, filter paper, and precipitate. Record this mass.

**16**. The precipitate and filter paper may be discarded in the waste basket. Wipe the marker off of the watch glass.

**Prelab Instrumentation and Procedure Questions**

**1.** Use the masses of sodium carbonate and calcium chloride to predict the mass of calcium carbonate that will form in your experiment.

**2.** If one more gram of sodium carbonate was used, how would it affect the amount of calcium carbonate that you *calculated* would form? Students may wish to test your answer by running the procedure again.

**3.** Which mass of the precipitate, the first or second, better represents the amount of dry precipitate collected? What mass of precipitate did you collect?

**4.** Is the mass you measured close to the expected mass you calculated based on stoichiometry in Question 1? What may be the reason(s) for any differences?

**5.** Would the mass of precipitate that you measured be larger or smaller if you did not wash the precipitate before drying it?

**6.** If the precipitate was weighed without drying, would you believe that you had started with 2 grams of calcium chloride? Explain.

**7.** Do you feel that the second weighing of your precipitate was dry? What experimental changes could be made to improve this portion of the procedure?

**Investigation helpful hints**

Students should determine the unknown quantity of Ca2+, as milligrams of CaCO3 per Liter of solution, in a water sample by designing their own procedure. This information should be presented in a typed form.

The materials listed are suggestions for students to use in their experiments. Quantities are suggestions based on having groups of 2–3 students. You may wish to substitute other materials. Should students wish to use additional materials that are not listed, check the availability of these materials in advance.

Students should consider what data are needed to obtain the concentration of the Ca2+ ions; this should be clearly exhibited in the step-by-step procedure and the creation of their data table. They should present a detailed, step-by-step procedure, a list of materials needed, and a data table for measurements and observations.

Procedures should indicate how the sodium carbonate solution will be used in excess, as well as how often and for how long the precipitate will be dried. Once Mr Clinch has approved the students’ proposed procedures, they may begin the data collection.

**Procedure**

Students will design their own procedures. Students should use the materials provided them but should check before deciding to use additional materials.

**Data Collection and Computation**

Students may either answer the following questions after completing the collection of all data.

**1.** How many grams of each precipitate were collected?

**2.** What is the hardness, in mg/L as CaCO3, of each water sample?

**3.** The water softener discussed here relies on precipitation softening, also known as ionexchange. Based on the reactants used (Na2CO3 and CaCl2), what ions would remain in the softened water that would be consumed by the homeowners? What could be some negative aspects of consuming these ions?

**4.** What other type of water softeners are available, besides the precipitation softening discussed here? Are there any advantages or disadvantages to using one of these, rather than an ion-exchange softener?

**Argumentation and Documentation**

Students will complete the Data Collection and Computation section on the side whiteboard and these data need to be installed in a data table in every student’s lab write-up.

When all groups are done, we will discuss the results. More than one group (preferably 3) must analyze the same water sample, and their results may be different. We will discuss what laboratory techniques may have been different between the groups.

After collecting and comparing the data, lab groups should submit a letter, directed to the Kenmore West Faculty. They should outline what was done, what was found, which community area has the lowest water hardness, and which water softener they would recommend.

Post Lab Assessment has been intentionally been removed from this lab. Students, at this point you have done enough work to earn credit.