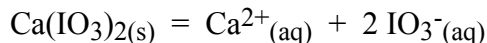


# Solubility Products

## Introduction

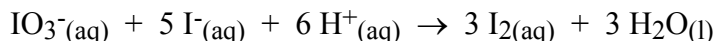
This experiment involves the determination of a solubility product constant. The calcium iodate chemical system to be analyzed is described by the reaction:



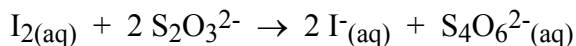
with a solubility product of:

$$K_{\text{sp}} = [\text{Ca}^{2+}] [\text{IO}_3^{-}]^2$$

In the first part of the experiment you will determine the solubility of calcium iodate in pure water. The solubility,  $s$ , of the calcium iodate will be equal to the concentration of the calcium ion since for every mole of calcium iodate that dissolves, one mole of calcium ion forms. Recall that the iodate ion concentration will be twice the calcium ion concentration in solution. Thus, if you can obtain the concentration of one ion you can calculate the concentration of the other ion. With the two concentrations you can easily calculate the solubility product constant. You shall determine the concentration of the iodate ion via what is known as an iodometric titration. In this process you will add excess iodide ion to solution that is known to contain iodate ion in the presence of acid. The iodate reacts with the iodide by the following reaction:



The  $\text{I}_2$  thus produced will then react via a titration with thiosulfate by the reaction:



It should be noted that the progress of this latter reaction can be followed because the iodine formed reacts with the excess iodide ion to form the triiodide ion,  $\text{I}_3^{-}$ . The presence of this species is easily observed by its reaction with starch to form a deep blue complex. Thus, in the presence of starch, the endpoint of this latter titration is when the deep blue color disappears. Once the concentration of the iodate has been determined you can easily calculate the concentration of the calcium ion and then the  $K_{\text{sp}}$  for the system.

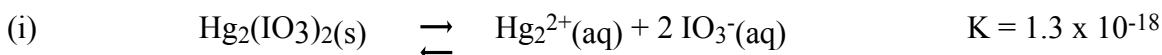
In the second part of this experiment you will be able to observe the "common ion effect". In this part of the experiment you will be given a saturated solution of calcium iodate in a 0.01 M potassium iodate solution. Once you determine the concentration of iodate by the method described above, you will be able to calculate the concentration of iodate from the dissolution of the calcium iodate and thus calculate the concentration of calcium ion in solution. Using the concentration of the two ions you will be able to calculate the solubility product constant for this system. By comparing the two parts you will note the dramatic effect that the iodate ion from the potassium iodate has on the solubility of calcium iodate.

Lastly as part of the data workup of this experiment, you will incorporate activity effects in the calculation of the solubility product from your data. The correct incorporation of activity effects makes the treatment of equilibria and equilibrium constants exactly rigorous.

You have discussed some of the effects of the polarity of water, including the effect that polarity can have on the solubility of solids. It should not be surprising to find that water interacts with different ions differently and that a more highly charged particle has a greater interaction with water molecules. The higher the charge on an ion in solution, the greater will be the interaction of the ion with the dipole of the water molecule and with other ions in the solution. These interactions can be significant enough that they cannot be ignored when salt concentrations exceed hundredth molar values.

Equilibrium constants are properly defined in terms of thermodynamic activity rather than concentration. The thermodynamic activity is a function of concentration, but is not necessarily equal to the concentration. However, it is true that in the limit of extremely dilute solutions the activity is equal to the concentration. Because the equilibrium constant expressions using concentrations in place of activity are rigorously correct in the limit of dilute concentration, are conceptually parallel with the use of activities, and because the results are useful, if not exactly correct, we commonly discuss equilibria and equilibrium constants using the concentrations. In this experiment, however, we will recognize that the true expressions are in terms of activities.

Based on the equilibrium constant of  $1.3 \times 10^{-18}$  for the dissolution of mercury(I) iodate, you would expect a saturated solution of the salt to be  $6.9 \times 10^{-7}$  M in mercury(I) ion.



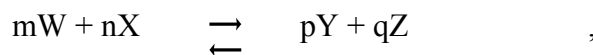
$$(ii) \quad K = [\text{Hg}_2^{2+}][\text{IO}_3^-]^2 = s(2s)^2$$

$$s = [\text{Hg}_2^{2+}] = 6.9 \times 10^{-7} \text{ M}$$

So far, when you have determined the effect of other dissolved ions on a specific equilibrium, you have only considered the common ion effect. Based on this reasoning, you would not predict that potassium nitrate in solution would have any effect on the solubility of mercury(I) iodate. However, if you were to saturate a 0.05 M potassium nitrate with mercury(I) iodate you would find that the solubility of the mercury(I) iodate has increased by about fifty percent. This turns out to be a general observation; any time you add an inert soluble salt to a solution of a sparingly soluble salt you will increase the solubility of the sparingly soluble salt.

The explanation for the observed increase in solubility is that the positively charged potassium ions can cluster around the negatively charged iodate ions, and the negatively charged nitrate ions cluster around the positively charged mercury(I) ions. When a mercury(I) ion comes close to an iodate ion surrounded by potassium ions, the positive charge on the potassium ions will repel the positive charge on the mercury(I) ion, preventing it from combining with the iodate ion and precipitating out of solution. Thus the mercury(I) iodate becomes more soluble.

The definition of the equilibrium constant represented in equation (ii) above does not take this phenomenon into account. Instead of looking only at the concentration of a species in solution, the **activity** of that species should be examined when equilibrium is considered. The activity of an ion includes both concentration and how susceptible the ion is to the kinds of effects described in the preceding paragraph. To incorporate these and other effects arising from molecular and ionic interactions in solution, we simply use the activity of the ion in place of the concentration in the equilibrium constant expression. The general way for incorporating activities into equilibrium constants for the general reaction:



is to form the equilibrium constant in the usual way, but employing activities where you have been using concentrations:

$$K = \frac{a_Y^p a_Z^q}{a_W^m a_X^n}.$$

Following this procedure for our example solubility problem, eq. (ii) becomes:

$$K = a_{\text{Hg}_2^{2+}} a_{\text{IO}_3^-}^2 \quad (\text{iii}).$$

A convenient way to quantitatively account for the molecular interaction part of the activity is to express the activity as the product of an activity coefficient times the concentration. For example the mercury iodate equilibrium requires the activities:

$$a_{\text{Hg}_2^{2+}} = \gamma_{\text{Hg}_2^{2+}} [\text{Hg}_2^{2+}]$$

$$a_{\text{IO}_3^-} = \gamma_{\text{IO}_3^-} [\text{IO}_3^-].$$

Where  $\gamma_{\text{Hg}_2^{2+}}$ , and  $\gamma_{\text{IO}_3^-}$  are the activity coefficients for  $\text{Hg}_2^{2+}$  and  $\text{IO}_3^-$ . Substituting these expressions into eq. (iii):

$$K = (\gamma_{\text{Hg}_2^{2+}} [\text{Hg}_2^{2+}]) (\gamma_{\text{IO}_3^-} [\text{IO}_3^-])^2 = \gamma_{\text{Hg}_2^{2+}} \gamma_{\text{IO}_3^-}^2 [\text{Hg}_2^{2+}] [\text{IO}_3^-]^2.$$

From this form we can see that expressing the equilibrium constant using concentrations alone is identical to assuming that the activity coefficients are equal to 1.0. This assumption is also called the ideal solution approximation.

Because it is impossible to get a solution containing just the cation or just the anion, it is impossible to experimentally determine  $\gamma_{\text{Hg}_2^{2+}}$  and  $\gamma_{\text{IO}_3^-}$  individually. Instead their product is replaced by  $\gamma_{\pm}$ , the mean ionic activity coefficient, raised to the power equal to the sum of the exponents of the individual ion activity coefficients.

$$K = \gamma_{\pm}^3 [\text{Hg}_2^{2+}] [\text{IO}_3^-]^2.$$

Since they account for molecular and ionic interactions, the values of activity coefficients change as the concentration of the solution changes. It has been found that a convenient quantity to use when expressing the functional dependence of the activity coefficients of ions on concentration is the ionic strength of the solution, which is defined by the expression:

$$\mu = \frac{1}{2} \sum_i c_i Z_i^2$$

where  $c_i$  is the concentration of the  $i$ th species and  $Z_i$  is its signed charge in multiples of the elementary charge (e.g.  $Z_{Hg_2^{2+}} = +2$  and  $Z_{IO_3^-} = -1$ ). This sum extends over all ions in solution.

In the example contrasting the solubility of mercury(I) iodate in pure water and in 0.05 M potassium nitrate it becomes very clear that the ionic strength of the solution in pure water is vastly different from the solution in 0.05 M potassium nitrate when we apply this definition.

In pure water

$$\mu = \frac{1}{2} (4[Hg_2^{2+}] + [IO_3^-]).$$

Since  $Hg_2^{2+}$  and  $IO_3^-$  are the only ions in solution. However in the solution containing potassium nitrate:

$$\mu = \frac{1}{2} (4[Hg_2^{2+}] + [IO_3^-] + [K^+] + [NO_3^-]).$$

Because mercury(I) iodate is so sparingly soluble, calculations will give the result that in pure water  $\mu = 0.0$  whereas in 0.05 M potassium nitrate  $\mu = 0.05$ .

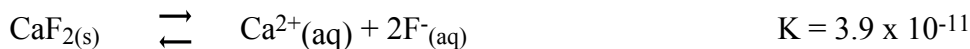
While it is impossible to experimentally determine the values of individual ion activity coefficients, various theoretical and empirical methods for consistently separating the observed mean ionic activity coefficients into individual ion coefficients have been developed. These methods are by no means perfect, but they often give much better results than the alternative very simple assumption that the solutions are ideal. **Table 1.** presents results for one such method of representing individual ion coefficients as a function of the ionic strength of the solution.

**Table 1. Activity Coefficients for Aqueous Solution at 25°**

Ion	Ionic Strength ( $\mu$ , M)					
	0.001	0.005	0.01	0.05	0.1	0.15
H <sup>+</sup>	0.967	0.933	0.914	0.86	0.83	0.81
Li <sup>+</sup>	0.965	0.929	0.907	0.835	0.80	0.77
Na <sup>+</sup> , IO <sub>3</sub> <sup>-</sup> , HCO <sub>3</sub> <sup>-</sup> , H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>	0.964	0.928	0.902	0.82	0.775	0.76
OH <sup>-</sup> , F <sup>-</sup> , SCN <sup>-</sup> , MnO <sub>4</sub> <sup>-</sup> , ClO <sub>4</sub> <sup>-</sup>	0.964	0.926	0.900	0.81	0.76	0.73
K <sup>+</sup> , Cl <sup>-</sup> , Br <sup>-</sup> , I <sup>-</sup> , CN <sup>-</sup> , NO <sub>3</sub> <sup>-</sup>	0.964	0.925	0.899	0.805	0.755	0.72
Rb <sup>+</sup> , Cs <sup>+</sup> , NH <sub>4</sub> <sup>+</sup> , Ag <sup>+</sup>	0.964	0.924	0.898	0.80	0.75	0.71
Mg <sup>2+</sup> , Be <sup>2+</sup>	0.872	0.755	0.69	0.52	0.45	0.41
Ca <sup>2+</sup> , Cu <sup>2+</sup> , Zn <sup>2+</sup> , Mn <sup>2+</sup>	0.870	0.749	0.675	0.485	0.405	0.36
Sr <sup>2+</sup> , Ba <sup>2+</sup> , Cd <sup>2+</sup> , Hg <sup>2+</sup> , S <sup>2-</sup>	0.868	0.744	0.67	0.465	0.38	0.33
Pb <sup>2+</sup> , CO <sub>3</sub> <sup>2-</sup> , SO <sub>3</sub> <sup>2-</sup>	0.867	0.742	0.665	0.455	0.37	0.31
Hg <sub>2</sub> <sup>2+</sup> , SO <sub>4</sub> <sup>2-</sup> , CrO <sub>4</sub> <sup>2-</sup> , HPO <sub>4</sub> <sup>2-</sup>	0.867	0.740	0.660	0.445	0.355	0.30
Al <sup>3+</sup> , Fe <sup>3+</sup> , Cr <sup>3+</sup>	0.738	0.54	0.445	0.245	0.18	0.15
PO <sub>4</sub> <sup>3-</sup>	0.725	0.505	0.395	0.16	0.095	0.066
Sn <sup>4+</sup>	0.588	0.35	0.255	0.10	0.065	0.048

Source: Data from J. Kielland, *J. Amer. Chem. Soc.*, **59**, 1675 (1937)

As an example of how to use activities, here is a calculation of the concentration of calcium ion in a 0.0125 M solution of magnesium sulfate MgSO<sub>4</sub> saturated with calcium fluoride CaF<sub>2</sub>. The concentration of calcium is going to depend on how much calcium fluoride dissolves, so the chemical equilibrium and initial set-up of interest is



Initial:	solid	0	0
Final	solid	x	2x

$$K = (a_{\text{Ca}^{2+}}) (a_{\text{F}^{-}})^2 = (\gamma_{\text{Ca}^{2+}} [\text{Ca}^{2+}]) (\gamma_{\text{F}^{-}}^2 [\text{F}^{-}]^2)$$

$$= (\gamma_{\text{Ca}^{2+}} [x]) (\gamma_{\text{F}^{-}}^2 [2x]^2) = 4\gamma_{\text{Ca}^{2+}} \gamma_{\text{F}^{-}}^2 x^3$$

In order to look up the activity coefficients in the table, it is necessary to know the ionic strength of the solution. The ionic strength is due to the dissolved magnesium sulfate and the dissolved calcium fluoride. Since the equilibrium constant for the dissolution of calcium fluoride is quite small, assume that, as in the earlier ionic strength calculation, its contribution will be negligible, and the only ions that need to be considered are the magnesium and sulfate ions. This gives an ionic strength of 0.050 M:

$$\mu = 1/2([0.0125] * 2^2 + [0.0125] * (-2)^2) = 0.050 \text{ M}$$

Using this value for the ionic strength, the activity coefficients for calcium and fluoride ions are 0.485 and 0.81 respectively. Plugging into the equilibrium constant equation and solving for x gives

$$3.9 \times 10^{-11} = [x]0.485[2x]^20.81^2$$

$$x = [\text{Ca}^{2+}] = 3.1 \times 10^{-4} \text{ M}$$

If you had neglected the activity of the ions in solution you would have calculated the calcium ion concentration to be  $2.1 \times 10^{-4} \text{ M}$ . This is a thirty-two percent error.

In this experiment you will examine the effect of activities in determining an equilibrium constant, the solubility product. You will do a calculation similar to the example given, but you will determine the concentrations of the species in solution, and you will use these to calculate the solubility product both with and without including activity effects. Your solutions are not likely to have an ionic strength exactly equal to one of those given in the table. While more sophisticated interpolation between values in the table are possible, it is sufficient for this experiment to simply use the tabulated value for that ionic strength that is closest to the value you calculate for your solution of interest. If your solution has an ionic strength exactly midway between two tabulated values, then use the value for the lower ionic strength.

**To prepare for this experiment review Chapter 19 of the text, Petrucci et. al., 8<sup>th</sup> ed.**

**Safety: Remember to always wear gloves when handling all solids. WEAR YOUR GOGGLES!**

**Work in pairs on this experiment.**

**Each student must collect data and submit a separate report.**

The actual data analyses and the written reports must be done entirely independently of your lab partner or other students. Make sure that you avoid unauthorized collaboration and plagiarism. All suspected violations of the Code of Academic Conduct will be referred to Student Judicial Affairs.

### **Experimental Procedure**

#### **Part I. Preparation of a Potassium Iodate Solution**

In this procedure you will prepare a standard potassium iodate solution for use in standardizing a sodium thiosulfate solution.

1. Using a volumetric flask, accurately prepare 250 mL of a solution approximately 0.01 M  $\text{KIO}_3$ . Note that this will need to be made as accurately as possible since this solution will serve as the primary standard for the experiment. The solid will take a few moments to dissolve. While you are waiting, go on to Part II.

#### **Part II. Preparation of a Sodium Thiosulfate Solution**

In this procedure you will prepare and standardize a sodium thiosulfate solution. You will use the standard  $\text{KIO}_3$  solution prepared in Part I to standardize the sodium thiosulfate solution. The sodium thiosulfate solution will then become a secondary standard.

1. Using your clean 1 L plastic bottle, prepare about 500 mL of a 0.05 M  $\text{Na}_2\text{S}_2\text{O}_3$  solution. You should do this using the 1.0 M  $\text{Na}_2\text{S}_2\text{O}_3$  stock solution provided. Some calculations are required here.

#### **Part III. Standardization of the Sodium Thiosulfate Solution**

Standardize the sodium thiosulfate solution with the potassium iodate solution you have prepared. You should use the iodate-to-iodine and iodine-to-iodide reactions given in the introduction of the experiment. You will have to plan how to best perform this experiment. Here are some tips:

1. For the first reaction you will need excess potassium iodide and hydrochloric acid. If you were using a 10 mL sample of the standard potassium iodate solution it would

require about 2 g of potassium iodide and about 3 mL of 6 M hydrochloric acid. Begin by dissolving the potassium iodide in about 50 mL of water in your titration flask. Next, accurately add 10.0 mL of potassium iodate using volumetric glassware. Finally, add the hydrochloric acid. The solution will turn brown due to the formation of iodine.

2. Now immediately titrate the solution with the sodium thiosulfate solution until the solution in the titration flask becomes a light brown color indicating only a limited amount of iodine present. At this point add 1 mL of the starch indicator. The starch will react with the iodine to form a complex that is dark blue in color. Continue the titration until the dark blue color just disappears.
3. Do at least three acceptable titrations so that you can calculate a meaningful average sodium thiosulfate concentration, standard deviation and 90% confidence limit for your data. An acceptable trial is one that passes the Q-test.

### **Titration Tips**

- Do not use assembly line techniques when preparing flasks for titration; prepare a flask, titrate it, and then prepare the next flask. If you do not begin the titration immediately iodine may crystallize out of solution and your titration results will be inaccurate.
- Do not add the starch indicator until the brown solution has lightened to light brown or yellow.
- Do not waste chemicals. You should be using no more than a couple of grams of potassium iodide and a few milliliters of acid in each titration.
- Do not waste time. Only your limiting reagent needs to be measured out with volumetric glassware. The other reagents in the iodate-to-iodine reaction can be measured out reasonably roughly without affecting your results.
- Do not contaminate your reagents. If the solutions in the bottles are brown they have been contaminated. Return these bottles to the stockroom.

### **Part IV. Solubility and Solubility Product from a Saturated Calcium Iodate Solution**

Determine the iodate concentration in a saturated solution of calcium iodate using your standard sodium thiosulfate solution. You should use the same procedure you developed for the thiosulfate standardization, but this time you will use a saturated calcium iodate solution instead of a potassium iodate solution when performing the first reaction.

The saturated calcium iodate solution will be provided by the stockroom. Make sure you read the labels on the bottles provided carefully; there are two calcium iodate solutions, and for this part of the experiment you are interested in the solution that has only calcium

iodate dissolved in water. Pour out the required amount carefully so that you do not disturb the solid  $\text{Ca}(\text{IO}_3)_2$  settled at the bottom of the bottle. **You do NOT want any of the solid  $\text{Ca}(\text{IO}_3)_2$  in your Erlenmeyer flask.**

You may only have time for a single titration for this part of the experiment. Do one titration then move on to Part V. If there is time remaining in the lab come back and repeat this titration.

### **Part V. Solubility and $K_{sp}$ from a Saturated Calcium Iodate Solution in 0.010 M $\text{KIO}_3$ .**

Determine the iodate concentration in a saturated solution of calcium iodate in 0.010 M  $\text{KIO}_3$  using standard sodium thiosulfate solution. This saturated solution is prepared by dissolving enough potassium iodate in water to make it 0.010 M in iodate and then saturating the solution with calcium iodate. This solution will also be provided by the stockroom. Again, read the label carefully; do not confuse this solution with the saturated calcium iodate solution you used in the previous part of this experiment. Pour out the required amount carefully so that you do not disturb the solid  $\text{Ca}(\text{IO}_3)_2$  settled at the bottom of the bottle. **You do NOT want any of the solid  $\text{Ca}(\text{IO}_3)_2$  in your Erlenmeyer flask.**

You only have time to do a single titration for this part of the experiment. Do one titration then go back to Part IV. After you have done a second titration for Part IV do a second titration for this part of the experiment if time permits.

Clean Up: All solutions can go down the drain. Clean and return all burets to the proper location at the back of the room. Clean up your glassware and your work area and complete any other cleanup tasks to which you have been assigned.

### **Data Analysis**

#### **Part I.**

1. What was the mass of  $\text{KIO}_3$  that you dissolved in 250.0 mL of de-ionized water to make your primary standard solution?
2. What was the resulting molarity of your primary standard solution of  $\text{KIO}_3$ ?

#### **Part II.**

3. What volume of 1 M  $\text{Na}_2\text{S}_2\text{O}_3$  stock solution did you use to prepare 500 mL of 0.05 M  $\text{Na}_2\text{S}_2\text{O}_3$ ?

#### **Part III.**

4. What is the stoichiometric factor, that is the number of moles of  $\text{Na}_2\text{S}_2\text{O}_3$  reacting with one mole of  $\text{KIO}_3$ ?
5. For each of your three trials, what volume of  $\text{Na}_2\text{S}_2\text{O}_3$  was required to reach the endpoint?

6. What is the molarity of the  $\text{Na}_2\text{SO}_3$  that you calculate for each of your three trials in the same order in which you entered the volumes?
7. What is the average molarity and the standard deviation of the  $\text{Na}_2\text{SO}_3$  solution based on your three trials?

**Part IV.**

8. How many trials were you able to complete for the determination of  $\text{IO}_3^-$  in the saturated solution of  $\text{Ca}(\text{IO}_3)_2$  in pure water solvent?
9. What volume of the saturated solution of  $\text{Ca}(\text{IO}_3)_2$  in pure water did you use as a sample for titration with  $\text{Na}_2\text{SO}_3$ ?
10. For each of the trials you performed, what volume of standardized  $\text{Na}_2\text{SO}_3$  was required to reach the endpoint?
11. For each of the trials you performed, how many moles of  $\text{IO}_3^-$  were present? If you performed multiple trials, what was the average number of moles of  $\text{IO}_3^-$  present?
12. Based on the number of moles of  $\text{IO}_3^-$  present in your sample(s) and the volume of that sample, what is the solubility of  $\text{Ca}(\text{IO}_3)_2$  in the saturated solution in pure water?

**Part V.**

You will now calculate the concentration of the iodate present in your final solution. When you do this calculation, make sure that you account for the fact that there is 0.010 M iodate present that does not come from the dissolved  $\text{Ca}(\text{IO}_3)_2$ .

13. How many trials were you able to complete for the determination of  $\text{IO}_3^-$  in the saturated solution of  $\text{Ca}(\text{IO}_3)_2$  in the solution containing 0.01M potassium iodate  $\text{KIO}_3$ ?
14. What volume of the saturated solution of  $\text{Ca}(\text{IO}_3)_2$  in 0.01 M  $\text{KIO}_3$  did you use as a sample for titration with  $\text{Na}_2\text{SO}_3$ ?
15. For each of the trials you performed, what volume of standardized  $\text{Na}_2\text{SO}_3$  was required to reach the endpoint?
16. For each of the trials you performed, how many moles of  $\text{IO}_3^-$  were present? If you performed multiple trials, what was the average number of moles of  $\text{IO}_3^-$  present?
17. Based on the number of moles of  $\text{IO}_3^-$  present in your sample(s) and the volume of that sample, what is the solubility of  $\text{Ca}(\text{IO}_3)_2$  in the saturated solution in 0.01 M  $\text{KIO}_3$ ?

### Comparison of solubility values

18. Examine the values you have obtained for the solubility of  $\text{Ca}(\text{IO}_3)_2$  in pure water and in 0.01 M  $\text{KIO}_3$ . Is the value of solubility significantly different in pure water and in 0.01 M  $\text{KIO}_3$ ? Explain any difference you may observe.

### Calculating $K_{\text{sp}}$ based on concentration

19. Using the concentration of  $\text{Ca}(\text{IO}_3)_2$  that you determined in the saturated solution in pure water, what is the value of  $K_{\text{sp}}$  that you calculate using the expression in concentrations alone?
20. Using the concentration of  $\text{Ca}(\text{IO}_3)_2$  that you determined in the saturated solution in 0.01 M  $\text{KIO}_3$ , calculate the value of  $K_{\text{sp}}$  that you calculate using the expression in concentrations alone?

### Calculating $K_{\text{sp}}$ based on activity

21. What is the ionic strength of the saturated solution of  $\text{Ca}(\text{IO}_3)_2$  in pure water?
22. Using the activity coefficients from **TABLE I**, appropriate for the ionic strength of the saturated  $\text{Ca}(\text{IO}_3)_2$  in pure water, calculate  $K_{\text{sp}}$  using activities?
23. What is the ionic strength of the saturated solution of  $\text{Ca}(\text{IO}_3)_2$  in 0.01 M  $\text{KIO}_3$ ?
24. Using the activity coefficients from **TABLE I**, appropriate for the ionic strength of the saturated  $\text{Ca}(\text{IO}_3)_2$  in 0.01 M  $\text{KIO}_3$  solution, calculate  $K_{\text{sp}}$  using activities?

### Comparison of $K_{\text{sp}}$ values

25. Examine the values you have obtained for  $K_{\text{sp}}$  for  $\text{Ca}(\text{IO}_3)_2$  in pure water and in 0.01 M  $\text{KIO}_3$  both recognizing the presence of activity effects and based upon concentrations alone. Is the value of  $K_{\text{sp}}$  significantly different in pure water and in 0.01 M  $\text{KIO}_3$ ? Is any difference you calculate affected by the recognition of activity effects?

**Conclusion** Reflect on the experimental procedures you have undertaken and the possible sources of error. Then write a summary paragraph comparing your results, commenting upon the common ion effect and its influence on solubility, and commenting upon activity effects and their influence on the  $K_{\text{sp}}$  value.

