

The Effect of Salts and Nonelectrolytes on the Solubility of Potassium Bitartrate

An Introductory Chemistry Discovery Experiment

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Solubility, the common ion effect and solubility product constants are important topics in introductory as well as advanced chemistry courses. In the last 15 years, a number of solubility product experiments have been published in this *Journal* (1–11). Although some of these deal with the effect of ionic strength on solubility, only a few deal with the common ion effect. None are what can be considered discovery labs.

The experiments discussed here are an extension of a Chemical Education Resources Modular Laboratory experiment that I developed recently (12). These experiments involve an examination of the effects of the salts NaCl, KCl, and MgSO_4 as well as the nonelectrolyte glucose on the solubility of potassium bitartrate in water. The chemicals used in these experiments are inexpensive and nontoxic and present no disposal problems. Because the solubilities are determined by titration of the bitartrate ion, the experiments are suitable for high school or college introductory chemistry courses. They provide excellent experiences for students prior to their learning about solubility equilibria in the classroom. The experiments are best done just before solubility product equilibrium is discussed in class so that the lab will be a true discovery experience.

Procedure

To obtain sufficient data in a reasonable amount of time, this experiment is best done as a class project with each pair of students being assigned part of the entire experiment. In the first part of the experiment, the effect of KCl on the solubility of potassium acid tartrate (KHT) is studied. Following this,

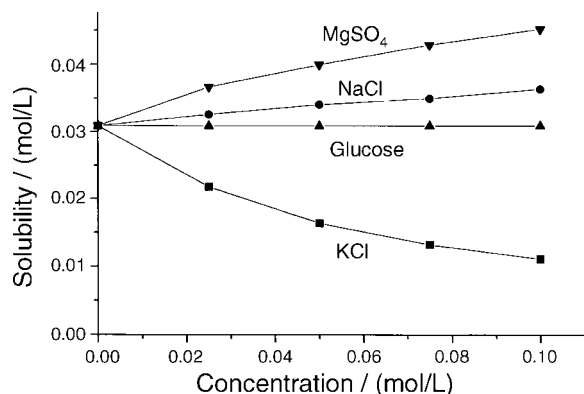


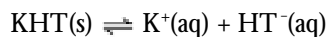
Figure 1. The effect of electrolytes and the nonelectrolyte glucose on the solubility of potassium bitartrate in aqueous solution. The solutions were saturated at 22.5 °C.

the effects of NaCl, MgSO_4 , and glucose on the solubility are investigated.

The students are provided with 0.10 M stock solutions of KCl, NaCl, MgSO_4 , and glucose. They are directed to prepare 200 mL of aqueous solutions with concentrations ranging between 0.0 M and 0.10 M. The concentrations can vary by increments of 0.01, 0.02, or 0.025 M depending on the number of lab groups doing the experiment. Approximately 2 g of KHT is added to each solution and the solutions are stirred for at least 20 minutes to achieve saturation. A constant-temperature bath is desirable but not essential. After the 20-minute saturation period, the solutions are filtered using qualitative medium-fast filter paper. Fifty-milliliter portions of the filtered supernatant solutions are titrated with standard NaOH of about 0.05 M concentration to the phenolphthalein end point. At least two titrations should be performed for each supernatant solution. The changes in solubility with concentration of the electrolyte are significant and students will easily reproduce the trends presented below.

Results and Discussion

When an aqueous solution is saturated with potassium bitartrate (KHT), the following equilibrium is established:



Because the only source of $\text{HT}^-(\text{aq})$ is from the dissolving of the KHT(s), the solubility of KHT(s) can be determined from titration results. The concentration of the HT^- is equal to the moles of KHT that dissolved per liter which is the solubility.

The results of these studies (Fig. 1) indicate that the solubility of KHT is unaffected by the presence of glucose, shows a slight increase with increasing NaCl concentration, and a larger increase with MgSO_4 concentration, and decreases significantly with increasing KCl concentration. The solutions in these studies were saturated at a temperature of 22.5 °C.

The decrease in solubility with KCl can be explained in terms of the common ion effect. When KHT(s) dissolves, $\text{K}^+(\text{aq})$ and $\text{HT}^-(\text{aq})$ ions are produced. Equilibrium is achieved when the rate at which ions from the solid phase enter the solution equals the rate at which ions in solution return to the solid phase. Because the latter rate increases with increasing concentrations of $\text{K}^+(\text{aq})$ and $\text{HT}^-(\text{aq})$ ions, the amount of KHT needed to dissolve to achieve this equilibrium state decreases with the amount of $\text{K}^+(\text{aq})$ initially present.

The lack of an effect of glucose on the solubility of KHT contrasts with the observed increases in solubility with increasing NaCl and MgSO_4 concentrations. If this experiment is used in a physical chemistry or other advanced course, the increases in solubility with NaCl and MgSO_4 can be explained

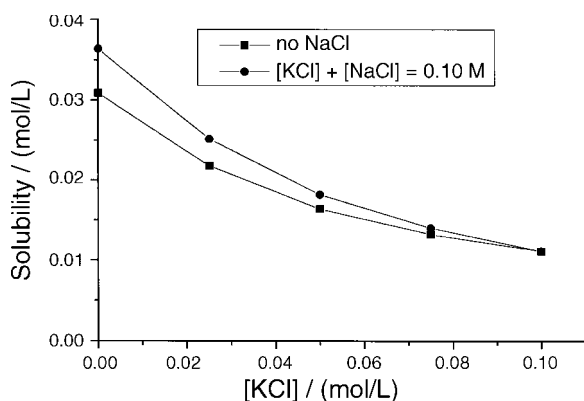


Figure 2. The effect of KCl and KCl–NaCl solutions on the solubility of potassium bitartrate in aqueous solution. The KCl–NaCl mixtures all have a constant ionic strength of 0.10 M. The solutions were saturated at 22.5 °C.

in terms of Debye–Hückel theory. However, on an introductory chemistry level, it is sufficient to say that the presence of non-common ions has a stabilizing effect on the ions involved in the equilibrium. This occurs because the K^+ and HT^- ions become surrounded by an ionic atmosphere that contains a slight excess of ions of opposite charge. The stabilization increases with increasing concentration of ions in solution (ionic strength). On a very basic level, one can say that this stabilization of the ions in solution has the effect of shifting the solubility equilibrium to the right. Alternatively, students can be told that the slight excess of ions of opposite charge around the K^+ and HT^- ions lowers their effective charge and reduces the force of attraction that they have for each other. This charge screening has the effect of reducing the tendency of the K^+ and the HT^- to aggregate. Thus, to achieve equilibrium in the presence of added electrolyte, higher concentrations of K^+ and HT^- must be present. Consistent with this explanation is the observation that $MgSO_4$ has a greater effect on the solubility of KHT than does NaCl.

After the concept of solubility product constant is discussed in class, the data obtained from these experiments can be used to calculate the apparent solubility product constant ($K_{sp} = [K^+][HT^-]$). When this is done, students will discover that the solubility product “constant” is not constant. In fact,

it increases from 9.5×10^{-4} to 1.3×10^{-3} as the NaCl concentration varies from 0.0 to 0.10 M, and it increases from 9.5×10^{-4} to 2.0×10^{-3} as the $MgSO_4$ concentration varies over the same concentration range. It is also found that K_{sp} increases from 9.5×10^{-4} to 1.2×10^{-3} as the KCl concentration varies from 0.0 to 0.10 M.

In order to obtain a constant K_{sp} , the experiment should be repeated by saturating solutions containing mixtures of KCl and NaCl so that $[NaCl] + [KCl] = 0.10$ M. The results of this experiment are shown in Figure 2. When this is done, students are gratified to discover that the apparent K_{sp} is remarkably constant ($1.2 \times 10^{-3} \pm 0.1 \times 10^{-3}$) over the entire range of KCl concentrations.

Obviously, it would be rather time consuming to have the students do all the experiments described here. The entire study can be completed by assigning a different solution mixture to each pair of students. Alternatively, selected portions of the experiment can be performed. If it is desired to avoid the issue of the effect of ionic strength on the solubility and K_{sp} , the solubilities of KHT in constant-ionic-strength mixtures of KCl and NaCl can be compared with those in NaCl–NaBr mixtures. The solubility will decrease with increasing concentrations of KCl in the KCl–NaCl solutions but will be independent of the composition of the NaCl–NaBr mixtures. As long as the ionic strength is kept constant, the apparent K_{sp} will be constant. The consistency of the experimental K_{sp} values when the experiment is performed this way is exceptional.

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