

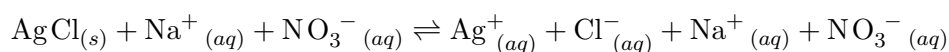
Chemistry 12 IB Spectator Ions and Solubility Equilibrium

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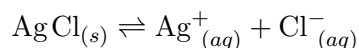
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1 Introduction

While performing chemical reactions, we usually neglect the effects of spectator ions on solubility equilibria as they are assumed to be non-participants in the reaction in question. For example, in the solubility equilibrium of silver chloride, addition of sodium nitrate does not disturb the equilibrium and thus the sodium and nitrate ions effectively “cancel” themselves out in the net ionic reaction.



which may be written as



The sodium and nitrate ions are thus called *spectator ions* as they do not affect the solubility of AgCl.

In reality however, it may be possible that these spectator ions do have an effect on solubility, but their effect is so insignificant that they can be virtually ignored. This lab aims to investigate this change in solubility, if any, and quantitatively determine the kind of relationship that exists between the amount of spectator ions present in the solution and the corresponding shift in solubility.

2 Planning A

2.1 Aim

To quantitatively determine the relationship between spectator ions and solubility equilibria by measuring K_{sp} values of solutions.

2.2 Hypotheses

We believe that spectator ions indeed have an effect on solubility equilibrium. However, this effect is to be so small that we can conveniently assume them to be true spectators in the two-way equilibrium reaction.

Although we know that spectator ions will disturb the equilibrium, to be able to hypothesize exactly what this effect will be is difficult as there are many factors to be considered. For example, the effect may depend on the solid being examined, its raw solubility and perhaps its acid-base properties. It is also difficult to predict in which direction the equilibrium will shift. If the presence of spectator ions result in the equilibrium shifting left, less ions would be produced and the solubility of the solid would decrease. On the contrary, if the spectator ions caused the equilibrium to shift right, more ions would be produced causing the solubility of the solid to increase.

2.3 Variables

Independent – Concentration of NaCl in solution

Dependent – K_{sp} value of solution

Controlled – Concentration of $\text{Ca}(\text{OH})_2$

3 Planning B

3.1 Materials Required

1. Hydrochloric acid (HCl) stock solution
2. Calcium hydroxide powder
3. Sodium bicarbonate (NaHCO_3) powder
4. Buret, pipet, beakers, flasks, suction bulb, 1 L bottle
5. Chemical indicator (bromcresol green)
6. Electronic Balance

3.2 Procedure

Pre-lab procedures

1. Prepare 25 mL solutions of fixed concentrations of $\text{Ca}(\text{OH})_2$ and varying concentrations of NaCl, and set aside for a couple of days to allow the solutions to reach equilibrium. Make sure there is at least one solution with no NaCl at all ($[\text{NaCl}] = 0$) that will act as a control.

Part I

1. Prepare a 1 L dilute solution of HCl and record its approximate concentration. Pour this solution into a bottle to be standardized in the next step.
2. Prepare a solution of NaHCO_3 of known concentration and pipet 10 mL into a beaker.

3. Prepare the buret by rinsing it thoroughly. Pour HCl into the buret and record its initial volume.
4. Add a few drops of indicator to the beaker and begin titrating until the expected colour change is observed. Record the final volume of HCl from the buret.
5. Repeat the above steps to obtain an average, and calculate the exact concentration of the now standardized HCl.

Part II

1. Pipet 10 mL from the saturated Ca(OH)_2 solution (containing no spectator ions) into a beaker, and add a few drops of indicator liquid.
2. Re-rinse the buret and fill it with the standardized HCl while recording its volume.
3. Titrate the Ca(OH)_2 with the HCl, and find out the concentration of OH^- by means of stoichiometric calculations. From this data, calculate the K_{sp} of the Ca(OH)_2 saturated solution.

Part III

1. Repeat the above part for each of the prepared saturated solutions and calculate corresponding K_{sp} values for each concentration of NaCl. Tabulate your findings.

4 Data Collection

Attached

5 Data Analysis

Attached

6 Evaluation

The following table summarizes our data analysis:

[NaCl]	K_{sp}
0.000	3.36×10^{-3}
0.110	2.44×10^{-6}
0.210	2.92×10^{-6}
0.315	7.09×10^{-4}
0.410	8.20×10^{-4}

Table 1: [NaCl] and Ksp

The closest value obtained was 2.44×10^{-6} which had percentage error of

$$\frac{2.44 - 1.3}{1.3} \times 100 \% = 87.7 \%$$

Obviously, this error is enormous and the values thus cannot be considered useful. However, even though the obtained values were inaccurate, it is possible to look for probable trends in the solubility constant of $\text{Ca}(\text{OH})_2$. It is obvious from the above table and the attached graph that the K_{sp} values follow no apparent trend. This was so because our titrations were highly erroneous and the values obtained cannot be considered correct. Thus, our results have absolutely no significance whatsoever as no useful conclusions can be drawn from them.

In order to identify trends, it would be best if this lab was repeated by improving the procedure and by keeping in mind sources of error.

6.1 Errors

Major sources for error in our experiment were:

1. **Inaccurate titrations.** Most of the titrations conducted were highly inaccurate as we were unable to tell when the end-point had been reached. Although we stopped at just almost the same colour (green), the colour change was so gradual that it was possible to add many more drops of the titrant before observing the expected colour.
2. **Unclean burets and pipets.** Because we had overlooked cleansing our burets and pipets thoroughly before using them, it is possible that some chemicals already existing in these apparatuses could have reacted with our acid. Even a slight amount of chemical (especially a base) in the buret would have had devastating effects as the concentration of our acid was relatively small. Any base already existing in the buret could have neutralized the acid leading us to unknowingly conduct a titration with water as the titrant.
3. **Excess indicator.** Since the indicator bromocresol green is itself a weak acid, it may have reacted with the base in the beaker and set up its own equilibrium even before the titration began. This becomes all-the-more important since the concentration and volume of our analyte in the beaker were considerably small.

6.2 Procedural Improvements

1. Use a pH meter to characterize end-point. This would help reduce our dependence on an indicator liquid.
 2. Add extra solid into each of the test-tubes so that more solid could be dissolved in the event that the equilibrium was shifted to the right due to the addition of spectator ions.
 3. Have more number of data points to spot any trends. The more number of data points, the more easier it is to identify inaccurate titrations.
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