# Topic Exploration Pack

# The Solubility Product

[Learning outcomes 2](#_Toc471889683)

[Introduction 2](#_Toc471889684)

[Suggested activities 3](#_Toc471889685)

[Activity 1: Finding the solubility product of calcium hydroxide 4](#_Toc471889687)

[Activity 2: Finding the solubility product of copper sulfate 7](#_Toc471889688)

[Activity 3: Calculations involving *K*sp and the importance of the solubility of calcium   
carbonate 10](#_Toc471889689)

[Learner Activity 13](#_Toc471889692)

[Learner Activity 1 Finding the solubility product of calcium hydroxide 13](#_Toc471889693)

[Learner Activity 2 Finding the solubility product of copper sulfate 17](#_Toc471889694)

[Learner Activity 3 Calculations involving *K*sp and the importance of the solubility of calcium carbonate 22](#_Toc471889695)

## Instructions and answers for teachers

These instructions cover the student activity section which can be found on [page 13](#_Learner_activity). This Topic Exploration Pack supports OCR A Level Chemistry B (Salters).

**When distributing the activity section to the students either as a printed copy or as a Word file you will need to remove the teacher instructions section.**

**123** **–** This activity offers an opportunity for maths skills development.

### Learning outcomes

O(h) the term *solubility product* for ionic compounds; solubility product calculations; techniques and procedures for determining solubility products

### Introduction

**Prior knowledge and student misconceptions**

As they approach this topic learners will be familiar with the idea of solubility being a property that can vary across a wide range of values, although they will not yet have met any numerical measurement of solubility.

Up to GCSE, learners are used to the idea that a salt is either soluble or insoluble, typically in water. This idea continues into the first year of the A Level course as they meet solubility rules and they assimilate knowledge concerning which salts are soluble or not in water. These rules are often applied in precipitation reactions, using the presence of white or coloured precipitates to test for the presence of ions, such as halide ions using acidified silver nitrate or sulfate ions using barium chloride.

The first time that they are introduced to the idea that a salt can be partially soluble is in the study of Group 2, where they meet the trends in solubility of the Group 2 sulfates and hydroxides. If done practically, the learners will see solutions of these compounds becoming increasingly cloudy as the solubility decreases.

Therefore, learners are now ready to make the next step. They appreciate the idea that solubility is a sliding scale with totally insoluble salts at one end and very soluble salts at the other. We are now taking this idea and making it numerical by asking the question: ‘How much of a certain salt can I dissolve in a fixed volume of water before the solution is saturated?’

In approaching the mathematics of the solubility product, it cannot be stressed too greatly that the learners are simply taking the idea of the equilibrium constant that they met in the first year and applying it to a new situation. Highlighting this link between areas of the specification at this point will put a great deal of learners more at ease and ready to explore this area.

**Key points in this topic**

The focus of this Topic Exploration Pack is the solubility product and the techniques by which we could find its value. We will also need to explore the calculations surrounding the value of the product.

The solubility product is the equilibrium constant for the reversible process of a salt dissolving in water.

salt(s)  salt(aq)

For this process to be at equilibrium, the rate at which the salt dissolves must equal the rate at which it is precipitating. This will only be seen in the situation in which there is an amount of the solid salt sitting at the bottom of a saturated solution of the salt. In the activities we explore how to set this up effectively.

For one of the compounds that we will explore, calcium hydroxide, the reaction is

Ca(OH)2(s)  Ca2+(aq) + 2OH–(aq)

The solubility product (equilibrium constant) for this process will be:

*K*sp  = [Ca2+(aq)][OH–(aq)]2

Many learners may now ask about the solid calcium hydroxide and why it is not included in the expression.

Solids do not appear in expressions for equilibrium constants (*K*). One justification is that the ‘concentration’ of a solid is a constant value, being equivalent to its density, so this constant value is assimilated within the constant value of *K*sp.

In the suggested activities, titration is used to find numerical values for *K*sp for two different salts. Then what is commonly known as the common ion effect is explored to see how further precipitation can be caused by adding different salts. This effect can be justified using Le Chatelier’s principle or, better, using the value of *K*sp.

### Suggested activities

#### Activity 1: Finding the solubility product of calcium hydroxide

Calcium hydroxide, often known as limewater when in solution, is a sparingly soluble compound and demonstrates the low solubility of the hydroxides towards the top of Group 2. In this activity, learners carry out a fairly straightforward acid–base titration to find the concentration of the solution of calcium hydroxide and then work through calculations to give a *K*sp value at room temperature.

#### Activity 2: Finding the solubility product of copper sulfate

This is a more challenging procedure and asks that the learners use two different techniques to find the solubility product of the salt: first using an iodine–thiosulfate titration and then using colorimetry.

The course does not explicitly cover colorimetry until Developing metals, but there is no objection to introducing it earlier. Learners will need to be familiar with the technique, especially if they wish to apply it to following rates of reaction. This provides a very suitable context to meet the method and understand the steps involved.

Learners are then invited to compare the two techniques and comment on the uncertainties associated with their final answers.

#### Activity 3: Calculations involving *K*sp and the importance of the solubility of calcium carbonate

Here the learners have a worksheet based activity in which they explore the mathematical demands of the expression for *K*sp and how this can be used to explain the role of carbonate rocks as carbon stores in the oceans.

### Additional teacher preparation

### Activity 1: Finding the solubility product of calcium hydroxide

Calcium hydroxide is not very soluble; in fact the value of its solubility product at 298 K is 5.5 × 10−6mol3dm−9. In this investigation, learners will use dilute hydrochloric acid to find the concentration of hydroxide ions in the saturated solution and then work back to find their value for *K*sp. They may then compare this value to a data book value and make a comment on the reasons for any difference.

As a starter, learners are asked to write the expression for *K*sp, which is

*K*sp = [Ca2+(aq)][OH–(aq)]2

Practical requirements

Chemicals

* solid Ca(OH)2 (around 0.5 g per pair/group) (DANGER; causes skin irritation and serious eye damage, may cause respiratory irritation)
* wash bottle of distilled water
* HC*l*(aq) of known concentration (around 0.05 mol dm−3) – 200 cm3 per pair/group
* phenolphthalein in dropper bottle (WARNING; highly flammable)

Apparatus and equipment

* 250 cm3 conical flasks ×2; 1 with stopper
* filter funnel and paper
* 25 cm3 pipette
* pipette filler
* small funnel
* burette, stand and clamp
* white tile

Health and safety

* Health and safety should always be considered by a centre before undertaking any practical work. A full risk assessment of any activity should always be undertaken.
* It is advisable to check the CLEAPSS website (<http://www.cleapss.org.uk>) in advance of undertaking the practical tasks.
* Learners should wear eye protection throughout.
* The calcium hydroxide solution should be treated as irritant.

The activity comes in three distinct parts: making the solution, titration and calculating *K*sp.

Task 1: Making a saturated solution of calcium hydroxide

This is actually more time-consuming than it originally appears. Learners would do well to complete this step at least 24 hours in advance to allow the equilibrium to establish.

At 298 K a saturated solution of calcium hydroxide will have a concentration around 0.018 mol dm−3, so the learners are going to measure out a large excess of the solid and attempt to dissolve as much as possible in around 150 cm3 of water.

To make around 150 cm3 of the saturated solution, learners should use about 0.5 g of calcium hydroxide. The exact volume of the solution does not matter as the learners will pipette out 25 cm3 samples to use in the titrations to find the concentration. Making about 150 cm3 should allow 1 trial and 4 accurate titrations, which should be enough.

A ‘quick review’ has been inserted in the learner sheet at this point to keep them focused on the purpose of the practical and to stop them from simply following orders. Suggested answers / considerations are given below.

1. A cloudy or milky solution is a sign of undissolved white precipitate; in this case it would be calcium hydroxide.
2. An equilibrium is a special situation for a reversible reaction in which the rates of the forward and reverse reactions are equal. Reaching this point may take some time and we should not think the equilibrium is suddenly reached just because all the required species are present.

Task 2: Titrating the saturated solution to find the concentration of hydroxide ions

The concentration of OH– ions in the saturated solution will be around 0.035 mol dm−3, so you might want to invite the learners to consider what concentration of hydrochloric acid would be suitable to give a reasonable value for the titre.

(A very small titre is not very good as it has a much higher percentage uncertainty associated with the reading, while a titre greater than the volume of the burette means that the learners have to stop before the meniscus drops below the last gradations, take the reading, refill, take down the new start point and then carry on. This is quite time-consuming, although it is good for the learners to see how this problem is dealt with.)

The learners are asked to produce the equation for the titration reaction.

Ca(OH)2(aq) + 2HC*l*(aq) → CaC*l*2(aq)+ 2H2O(l)

In this method, using HC*l* with a known concentration around 0.05 mol dm−3 would give a good titre if we use 25.0 cm3 of the saturated solution of calcium hydroxide.

The learners will now titrate to find the volume of acid required to neutralise the saturated calcium hydroxide solution. The learner sheet has a fairly explicit method, although competence in titration is expected.

This is a good opportunity to carry out formative assessment on the procedure of titration and to ensure that learners are still confident in the layout of a results table.

Task 3: Calculating the solubility product

In this final step the learners will go through the steps required to find a numerical value for the solubility product of calcium hydroxide. To find the concentration of the OH− ions, they must use the result of the titration.

The worksheet provides guidance as to the steps, although you may like to annotate the sheet to give more guidance depending on the needs of the learners.

To introduce a greater level of stretch and challenge for more able learners, simply do not give them any extra guidance in how they will use their titre figure to find the value of *K*sp.

Extra guidance could include:

* Use your mean titre to calculate the amount of acid required to react with the alkali in the conical flask.
* The acid and alkali reacted according to: H+(aq) + OH–(aq) → H2O(l)
* State the amount of OH– in the conical flask.
* If the volume of alkali in the flask was 25 cm3, calculate the concentration of OH– in the solution.
* To find the concentration of Ca2+ ions, use the equation for the dissolving process. 1 mol Ca(OH)2 dissolves to give 1 mol Ca2+ ions and 2 mol OH– ions.
* What is the ratio of the concentrations of Ca2+ to OH-?
* Use your concentration of OH– ions to calculate the concentration of Ca2+ in the solution.

Further questions explore why their final figure may differ from the data book value. Possible reasons for a difference could include:

* not doing the procedure at 298 K
* not allowing the solution to reach equilibrium, so the solution was not saturated
* poor filtration, allowing some solid through that would then dissolve as the titration proceeded
* distilled water used to make the solution contaminated by Ca2+ ions.

### Activity 2: Finding the solubility product of copper sulfate

This activity uses two different techniques to find the value of *K*sp for copper sulfate: iodine–thiosulfate titration and colorimetry. These techniques are formally covered in the course in Elements from the sea and Developing metals, respectively, but can be usefully reviewed or introduced here.

They are looking at the equilibrium established between solid copper sulfate and a saturated solution of the same salt, i.e.

CuSO4(s)  Cu2+(aq)+ SO42–(aq)

As an equilibrium is set up, we can write an expression for the solubility product, *K*sp:

*K*sp = [Cu2+(aq)][SO42–(aq)]

The learners now set out to find the numerical value of the solubility product at 298 K. The worksheet also asks about the solution at equilibrium. A good answer would be: ‘The solution must be saturated. This is because the rate of dissolving and the rate of precipitation is the same, which is a feature of equilibrium. The solid is still dissolving into the water, but because the solution is saturated, the solid is precipitating out at the same rate.’

Practical requirements - titration

Chemicals

* ×10 dilution of saturated CuSO4·5H2O(aq) (about 50 cm3 for each pair/group); preparation of the saturated solution is best done 24 hours in advance of the practical to ensure equilibrium has been reached; saturated CuSO4 is about 1.3 mol dm–3, so the ×10 dilution is about 0.13 mol dm–3 (WARNING; causes skin and eye irritation)
* KI(aq), approx. 1 mol dm–3 (about 75 cm3 for each pair/group)
* Na2S2O3(aq) of known concentration, approx. 0.1 mol dm–3 (about 100 cm3 for each pair/group)
* starch solution (about 10 cm3 for each pair/group)
* wash bottle of distilled water

Apparatus and equipment

* 10 cm3 pipette
* pipette filler
* 250 cm3 conical flask
* 25 cm3 measuring cylinder
* dropping pipette
* burette, stand and clamp
* white tile

Practical requirements – colorimetry

Chemicals

* ×10 dilution of saturated CuSO4·5H2O(aq) (a few cm3 for each pair/group) (WARNING; causes skin and eye irritation)
* CuSO4·5H2O(aq), 0.5 mol dm–3 (10–15 cm3 for each pair/group) (DANGER; causes skin irritation and serious eye damage)
* wash bottle of distilled water

Apparatus and equipment

* 10 cm3 measuring cylinder
* dropping pipette
* colorimeter
* cuvettes
* suitable filter

Health and safety

* Health and safety should always be considered by a centre before undertaking any practical work. A full risk assessment of any activity should always be undertaken.
* It is advisable to check the CLEAPSS website ([http://www.cleapss.org.uk](http://www.cleapss.org.uk/)) in advance of undertaking the practical tasks.
* Learners should wear eye protection throughout.

Teacher preparation for the practical

Copper sulfate is a fairly soluble salt; and is typically supplied as CuSO4·5H2O, with a molar mass of 249.7 g mol–1. The intent is to make a small volume of a saturated solution with undissolved solid in the beaker. This can then be filtered and then diluted by a factor of 10 to allow the practical work to be done without an excessive use of the chemicals and for use in the colorimetry.

1. Weigh out approximately 40 g of copper sulfate-5-water and attempt to dissolve it in about 100 cm3 of water. The copper sulfate is in excess, so the exact amounts do not matter.
2. Warm the solution gently to dissolve as much as possible.
3. Leave the solution to cool back to room temperature overnight to leave a deep blue solution and crystals at the bottom of the beaker.

This saturated solution is too concentrated for our techniques, so it will need to be diluted by a factor of 10. For use in the techniques described, learners will need around 60 cm3 of the diluted solution per group.

If equipment is available, you may want to consider doing a microscale titration to reduce use and disposal of copper sulfate. An excellent method for an iodine–thiosulfate titration can be found from the [Royal Society of Chemistry](http://www.rsc.org/learn-chemistry/resource/res00000539/measuring-the-amount-of-vitamin-c-in-fruit-drinks?cmpid=CMP00000611).

Task 1: analysis of the solution using iodine–thiosulfate titration

While the learners should have used this procedure in their first year of A Level study, they may need to be reminded of the steps involved. Once again, the learner worksheet provides guidance. In general terms, they will pipette 10 cm3 of the diluted solution that has already been made up and add an excess (15 cm3) of potassium iodide, which reacts with the copper to produce iodine:

2Cu2+(aq)+ 4I–(aq) → 2CuI(s) + I2(aq)

The iodine produced is titrated against sodium thiosulfate of known concentration:

I2(aq)+ 2S2O32–(aq)→ 2I–(aq) + S4O62–(aq)

Starch is used to highlight the end-point for this titration, producing a deep blue-black colour in the presence of iodine. From the titre value, we can find the amount of iodine produced in the first reaction and hence work out the amount of copper in the original 10 cm3 sample.

Once the learners have the concentration of Cu2+ in the original saturated solution, it is a small step to find the value of *K*sp from its expression.

Calculation steps to process the titration results:

1. Amount of Na2S2O3 = concentration × mean titre
2. Working through the equations shows that the amount of Cu2+ in the sample is the same as the amount of S2­O32–.
3. Concentration of diluted CuSO4(aq) = amount of Cu2+ / 10 cm3
4. Concentration of saturated solution = concentration of diluted solution × 10
5. The ions are present in a 1 : 1 ratio, so the concentration of SO42– is equal to the concentration of Cu2+.
6. *K*sp is the product of the answers to steps 4 and 5. Units are mol2 dm–6.

Task 2: Using colorimetry to find the value of the solubility product

This procedure is one that is worth taking time to examine to ensure that the learners fully understand the steps required. It is fairly lengthy, but important as it also appears in Developing metals when the learners look at colour in transition metal complexes.

Learners make up a series of known concentrations of copper sulfate solution and measure the absorbance of light through each solution in the colorimeter. This then produces a calibration curve relating concentration to absorbance. The solution of unknown concentration is then run through the colorimeter and the reading is converted to a concentration using the calibration curve already produced. As in Task 1, once the learners have a value for the concentration of Cu2+ ions in the solution, they can use this to find the value of the solubility product.

The learners are now invited to compare the results they obtained for *K*sp from the two different techniques. Some points for discussion could be:

* Did they get different results? How different were they?
* Can they consider potential sources of error and uncertainties in each method, and use this to assess which result is more likely to be accurate?
* How does the value from each technique compare with the databook value?

### Activity 3: Calculations involving *K*sp and the importance of the solubility of calcium carbonate

In this activity learners use expressions for the solubility product of salts to calculate amounts of substance in solutions or forming as precipitates. They will then expand these ideas to look at how the solubility product of calcium carbonate is important in the formation of carbonate rocks, which act as a store of carbon dioxide from the air. This can be set as a piece of work in class or as further practice of solubility calculations at home as an assignment.

Answers

Task 1

1. BeSO4(s)  Be2+(aq) + SO42–(aq)
2. *K*sp = [Be2+(aq)][SO42–(aq)]
3. (a) 391 g

(b) 391 / 105.1 = 3.72 mol dm–3

(c) The concentration of both ions is 3.72 mol dm–3.

(d) *K*sp = 3.72 × 3.72 = 13.8 mol2 dm–6

1. The solubility of the Group 2 sulfates decreases down the group, so the value of *K*sp would also be predicted to decrease going down the group.
2. MgSO4; solubility per dm3 = 351 g

351 / 120.4 = 2.92 mol dm–3

2.92 × 2.92 = 8.50 mol2 dm–6

CaSO4; solubility per dm3 = 2.55 g

2.55 / 136.2 = 0.0187 mol dm–3

0.0187 × 0.0187 = 3.51 × 10–4 mol2 dm–6

SrSO4; solubility per dm3 = 0.132 g

0.132 / 183.7 = 7.19 × 10–4 mol dm–3

7.19 × 10–4 × 7.19 × 10–4 = 5.16 × 10–7 mol2 dm–6

BaSO4; solubility per dm3 = 2.4 × 10–3 g

2.4 × 10–3 / 233.4 = 1.03 × 10–5 mol dm–3

1.03 × 10–5 × 1.03 × 10–5 = 1.1 × 10–10 mol2 dm–6

Task 2

1. Solubility falls at lower temperatures. So, as the solution cools it cannot hold as much solute. Salt will precipitate from the solution, becoming visible as a solid.
2. (a) *K*sp = [Sr2+][OH–]2; [OH–] = 2 × [Sr2+], therefore *K*sp = 4 × [Sr2+]3

At 90 °C: 18.8 = 4[Sr2+]3 ∴ [Sr2+]3 = 4.7 ∴ [Sr2+] = 1.68 mol dm–3

Therefore [Sr(OH)2] = 1.68 mol dm–3

mass dissolved per dm3 = 1.68 × 121.6 = 204 g

mass dissolved in 250 cm3 = 204 / 4 = 50.9 g

By the same route, the mass dissolved in 250 cm3 at 20 °C is 2.02 g.

(b) 50.9 g – 2.02 g = 48.9 g

Task 3

1. (a) Using amount = concentration × volume; *n*(BaC*l*) = 2.5 × 10–4 mol;   
   *n*(Na2SO4) = 2.5 × 10–3 mol

(b) New volume is 50 cm3. [Ba2+] = 5.0 × 10–3 mol dm–3; [SO42–] = 5.0 × 10–2 mol dm–3.

(c) [Ba2+][SO42–] = 5.0 × 10–3 × 5.0 × 10–2 = 2.5 × 10–4 mol2 dm–6

(d) The calculated value is much greater than the value for *K*sp. A precipitate will form. (Some learners may notice that this is the distinct white precipitate seen when barium chloride is used as a test for the presence of sulfate ions.)

1. *n*(A*l*C*l*3) = 5.0 × 10–6 mol; *n*(NaOH) = 5.0 × 10–5 mol

[A*l*3+] = 5.0 × 10–5 mol dm–3; [OH–] = 5.0 × 10–4 mol dm–3

[A*l*3+][OH–]3 = 5.0 × 10–5 × (5.0 × 10–4)3 = 6.25 × 10–15 mol4 dm–12

This is much greater than the value for *K*sp so a precipitate will form.

1. *n*(Pb(NO3)2) = 5.0 × 10–4 mol; *n*(NaC*l*) = 2.5 × 10–3 mol

[Pb2+] = 6.67 × 10–4 mol dm–3; [C*l*–] = 3.33 × 10–3 mol dm–3

[Pb2+][C*l*–]2 = 6.67 × 10–4 × (3.33 × 10–3)2 = 7.41 × 10–9 mol4 dm–12

This is much smaller than the value for *K*sp so no precipitate will form.

Task 4

1. CaCO3(s)  Ca2+(aq) + CO32–(aq)

*K*sp = [Ca2+][CO32–]

1. The value of *K*sp is very small, so calcium carbonate is barely soluble in sea water. The shells of living things will not dissolve in sea water. Calcium carbonate will easily precipitate from sea water to form shells.
2. Sodium (Na+) and chloride (C*l*–) ions do not appear in the *K*sp expression for calcium carbonate.
3. A lower temperature favours the reaction in the exothermic direction. So, calcium carbonate becomes *more* soluble at lower temperature, and therefore more soluble in the deep oceans.
4. CO2(aq) + H2O(l)  CO32–(aq) + 2H+(aq)

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# Topic Exploration Pack

# The Solubility Product

## Learner activity

### Learner Activity 1 Finding the solubility product of calcium hydroxide

In this activity you will find the solubility product of calcium hydroxide by

* making up a saturated solution
* carrying out a titration
* calculating the concentrations of the ions to substitute into the expression for *K*sp.

When excess solid calcium hydroxide is mixed with a saturated solution of the same compound, an equilibrium will establish:

Ca(OH)2(s)  Ca2+(aq) + 2OH–(aq)

Write an expression for the solubility product, *K*sp, for this equilibrium.

Calcium hydroxide is only sparingly soluble at room temperature (298 K), so the value of *K*sp will be very small.

Health and safety

* Wear eye protection.
* Solid calcium hydroxide causes skin irritation and serious eye damage, and may cause respiratory irritation. Take care to avoid skin contact.
* The calcium hydroxide solution should be treated as irritant.
* Phenolphthalein is flammable.

Task 1: Making a saturated solution of calcium hydroxide

To form the saturated solution required for this procedure, you will now attempt to dissolve an excess of calcium hydroxide in water. If this mixture is left to reach equilibrium, the solution will be saturated and we can analyse to find the concentrations of ions present.

1. Weigh out approximately 0.5 g of solid calcium hydroxide, Ca(OH)2, and place it in a conical flask with around 150 cm3 of water.
2. Swirl the flask to dissolve as much as you can; take care not to splash the solution out of the flask.
3. Stopper the flask and leave for the equilibrium to establish. The solution is best left overnight before moving on to the titration.

Quick review – discuss these questions in your group

1. Your mixture was cloudy or milky at the start. What does this suggest about the mixture at this point?
2. This topic is an example of an equilibrium. Suggest why it is necessary to leave the solution and solid for so long before analysis.

Task 2: Titrating the saturated solution to find the concentration of hydroxide ions

You are now ready to conduct an acid–base titration on the saturated solution. You will use a 25.0 cm3 sample of your saturated solution and titrate against hydrochloric acid of known concentration. You will need to check with your teacher to find the exact concentration of the acid being used in your experiment.

The equation for the acid–base reaction is

calcium hydroxide + hydrochloric acid → calcium chloride + water

Write the balanced equation for this reaction.

You will need this later in the analysis.

You are expected to carry out this titration to get an accurate titre by yourself, so this is a very simple set of instructions. You need to use your skills to get the most accurate value you can for the titre.

1. Filter the solution into a clean conical flask.
2. Prepare a suitable table in your notes for the results of the titration.
3. Safely pipette out 25.0 cm3 of the calcium hydroxide solution and place it in a fresh conical flask. Add a suitable indicator.
4. Fill the burette with the acid solution and titrate to find the end-point, repeating as necessary in order to get concordant results.

Task 3: Calculating the solubility product

To calculate the value of *K*sp for calcium hydroxide, you will need to analyse the results of the titration. You should be fairly confident about the steps in the calculation, so the guidance given below is quite sketchy.

Write down the expression for the *K*sp of calcium hydroxide that you generated earlier.

You can see that you will need to find the concentrations of both the OH– ions and the Ca2+ ions, and substitute these into the expression.

1. Using the titre from the titration, calculate the concentration of OH– ions in the solution.

2. Consider the equation for the dissolving process:

Ca(OH)2(s)  Ca2+(aq) + 2OH–(aq)

Calculate the concentration of the Ca2+ ions in the solution.

1. Substitute your calculated concentrations values into the expression for *K*sp to come up with a final answer. Don’t forget the units!
2. The actual data book value for the solubility product of calcium hydroxide is 5.5 × 10−6mol3 dm–9. How close was your answer?
3. Suggest any factors that you think may have contributed to a difference between your answer and the data book value.

### Learner Activity 2 Finding the solubility product of copper sulfate

In this activity you are going to use two different techniques to find the concentration of copper sulfate in a saturated solution: redox titration and colorimetry. This will allow you to calculate a value for the solubility constant of copper sulfate.

Your teacher has already made up a saturated solution of copper sulfate, but to make the experiments work more easily, the saturated solution has been diluted by a factor of 10.

Write the equation for the reversible process of dissolving copper sulfate.

Write an expression for the solubility constant, *K*sp, for this equilibrium.

Health and safety

* Wear eye protection.
* Treat all copper sulfate solutions as causing skin irritation and serious eye damage.

Task 1: Analysis of the solution using iodine–thiosulfate titration

You should have studied the iodine–thiosulfate titration in the first year of your A Level course. In this case, you are using it to find the concentration of the copper ions in the solution. You will use a 10 cm3 sample of your saturated solution, react this with copper(I) iodide solution, and titrate the resulting mixture against sodium thiosulfate solution of known concentration. You will need to check with your teacher to find the exact concentration of the thiosulfate solution being used in your experiment.

Here is a basic method. You are expected to use your knowledge of titrations to complete this activity to obtain an accurate final result.

1. Pipette out 10.0 cm3 of the copper sulfate solution and add it to a conical flask.
2. Add 15 cm3 of potassium iodide (1 mol dm–3) to the flask. This is a great excess, so do not worry about the exact volume. A white precipitate will form, which is copper(I) iodide, and the solution will turn brown due to the formation of iodine.

2Cu2+(aq) + 4I–(aq) → 2CuI(s) + I2(aq)

1. Fill the burette with the sodium thiosulfate solution.
2. Prepare a suitable results table and carry out the titration.

This is the reaction:

I2(aq) + 2S2O32–(aq) → 2I–(aq) + S4O62–(aq)

1. The end point is marked by the disappearance of the brown iodine colour. This is quite hard to see, so add 1 cm3 of starch solution as you approach end point. This turns blue-black in the presence of iodine and the disappearance of this colour is much easier to spot.
2. Repeat as necessary to obtain concordant results.

Calculation steps

1. Calculate the amount of sodium thiosulfate required to reach end point in the titration.
2. Now use the equations to work out the amount of Cu2+ in the original 10 cm3 sample.
3. Calculate the concentration of the diluted copper sulfate solution.
4. Work out the concentration of Cu2+ ions in the saturated solution.
5. The equation for the dissolving process is

CuSO4(s)  Cu2+(aq) + SO42–(aq)

State the concentration of SO42– ions in the saturated solution.

1. Substitute your values for the concentrations into the expression for *K*sp to come up with a final answer. Don’t forget the units!

Task 2: Using colorimetry to find the value of the solubility product

This task allows you to use colorimetry as another tool to find the concentrations of the ions in the solution.

Before you can start analysing your copper sulfate solution, you will need to produce a calibration curve to convert the readings on the colorimeter into concentrations in mol dm–3.

**Making a calibration curve**

1. Insert a suitable filter into the colorimeter. As the solution we are using is blue, the suitable filter is the complementary colour, orange.
2. First run the colorimeter with distilled water in a cuvette, as this will be used to ‘zero’ the meter.
3. Fill one of the cuvettes with 0.5 mol dm–3 copper sulfate solution.
4. Insert the cuvette into the colorimeter and take an absorbance reading.
5. Now you will need to repeat the readings for a series of diluted solutions of copper sulfate. Use the 0.5 mol dm–3 solution to make solutions of 0.4, 0.3, 0.2 and 0.1 mol dm–3. You only need a very small volume of each, but be as accurate as possible in the dilutions.
6. You will now have a table of results of absorbance for all the concentrations. Use this table to plot a graph of absorbance against concentration and complete a curve of best fit. This is your calibration curve!
7. You are now ready to analyse your ‘unknown’ solution in the experiment. Add a small volume to a cuvette and take the absorbance reading from the colorimeter. Find this value of absorbance on the calibration curve and read off the corresponding concentration of Cu2+ ions.

**Analysis**

You now have the concentration of Cu2+ ions in the diluted solution. The original saturated solution was diluted by a factor of 10, so what was the concentration of Cu2+ ions in the saturated solution?

The equation for the dissolving process is

CuSO4(s) Cu2+(aq) + SO42–(aq)

State the original concentration of the SO42– ions.

Use your expression for *K*sp, the solubility product of copper sulfate, and substitute your concentrations to get a final answer for *K*sp. Remember the units!

**Conclusion**

How do the final values for *K*sp from the two activities match up?

Are they significantly different?

What is the difference as a percentage of the actual values?

Can you suggest any reasons for the difference?

Which technique do you think has given the most accurate result? (An accurate result is one that is close to the true value.) Why?

### Learner Activity 3 Calculations involving *K*sp and the importance of the solubility of calcium carbonate

Task 1: Using solubilities to calculate a value of *K*sp for a partially soluble salt

The solubility of the Group 2 sulfates shows a trend as we go down the group from beryllium to barium.

Consider beryllium sulfate.

1. Write an equation for the reversible change when beryllium sulfate dissolves in water:
2. Write an expression for the solubility product for the above equilibrium.
3. The solubility of BeSO4 is 39.1 g per 100 cm3 of water at 293 K. This means that 39.1 g of the salt dissolves in 100 cm3 of water to give a saturated solution. Use this figure to calculate a value for the solubility product of BeSO4. Use the following steps as guidance.

(a) What mass of BeSO4 will dissolve in 1 dm3 of water?

(b) Convert this mass to amount in moles to find the solubility in mol dm–3.

(c) Thinking about the equation for the dissolving of BeSO4, what will be the concentration of the Be2+ and the SO42– ions in the solution?

(d) Substitute the concentrations into the expression for *K*sp to find a value. Don’t forget the units.

1. The solubility of the sulfates of the alkaline earth metals follows a trend as we go down the group from beryllium to barium. How would you predict the value of the solubility product to vary going down the group for the sulfates?
2. Now use the above guideline steps to find the value of *K*sp for all the Group 2 sulfates. You will need the following data.

| **Group 2 sulfate** | **Solubility / g per 100 cm3 water** |
| --- | --- |
| MgSO4 | 35.1 |
| CaSO4 | 0.255 |
| SrSO4 | 0.0132 |
| BaSO4 | 2.4 × 10–4 |

Task 2: Using solubility products to predict precipitation

The solubility of solids typically increases as temperature increases. So, if you make a saturated solution at a high temperature, then as the solubility drops when the solution cools, there will be a change called precipitation.

1. A saturated solution of a salt that behaves as described above is made at 90 oC.   
   What will be observed as the temperature of the solution drops to room temperature?

We can now use the changing values of *K*sp at different temperatures to work out the actual mass of solid salt that will precipitate out of solution as it cools.

1. At 90 °C strontium hydroxide, Sr(OH)2, has a Ksp of 18.8 mol3 dm–9. At 20 °C the value of Ksp is 1.18 × 10–3 mol3 dm–9.
2. What mass of Sr(OH)2 would dissolve in 250 cm3 of water at 90 °C and at 20 °C?

(b) What mass of strontium hydroxide would precipitate from 250 cm3 of the saturated solution if allowed to cool from 90 to 20 oC?

Task 3: the common ion effect

When we use *K*sp to complete a calculation, it can often be the case that the ions come from two different salts, not just one. However, the value of *K*sp at any one temperature must always be the same, so we need to think about the concentrations of ions from each source.

1. What would happen if we mixed 25 cm3 of 0.01 mol dm–3 BaC*l*2(aq) and 25 cm3 0.1 mol dm−3 Na2SO4(aq)?

*K*sp of BaSO4 at 298 K is 1.1 × 10–10 mol2 dm–6.

(a) First calculate the amount of each salt that is in the mixture.

b) Now calculate the concentration of the Ba2+ and SO42– ions in the solution. Remember that the total volume of the solution is the sum of the volumes that were added.

(c) Substitute the values for the concentration into the expression for *K*sp.

(d) How does this answer compare with the actual value of *K*sp at 298K? What do you now think will happen when these solutions are mixed?

1. Will a precipitate of aluminium hydroxide, A*l*(OH)3, form if we mix 50 cm3 of 1.0 × 10−4 mol dm–3 A*l*C*l*3(aq) and 50 cm3 of 1.0 × 10–3 mol dm–3 NaOH?

*K*sp of A*l*(OH)3 = 6.3 × 10–32 mol4 dm–12 at 298 K.

1. Will a precipitate of lead(II) chloride, PbC*l*2, form if we mix 500 cm3 of 0.001 mol dm–3 Pb(NO3)2 and 250 cm3 of 0.01 mol dm–3 NaC*l*?

*K*sp of PbC*l*2 = 1.6 × 10–5 mol3 dm–9 at 298 K.

Task 4: The solubility product of calcium carbonate

Calcium carbonate is a solid that forms a variety of sedimentary rocks, including chalk and limestone. These rocks were typically formed in shallow seas and have been lifted up by tectonic processes to give limestone landscapes.

In sea water, solid calcium carbonate is in equilibrium with a solution of its ions.

1. Write the equation for this equilibrium and the expression for *K*sp that goes with this reaction.
2. The value for *K*sp at 298 K is 5.0 × 10–9 mol2 dm–6. Sea creatures use the dissolved calcium carbonate to build shells. By discussing the value of *K*sp, explain why the calcium carbonate is a very useful building material for shells?
3. Sea water is a concentrated solution of many salts, mainly sodium chloride. Why does the presence of this salt not affect the solubility of calcium carbonate?
4. The process of dissolving calcium carbonate is exothermic. Describe and explain what will happen to the solid calcium carbonate in deep oceans, where the temperature is much lower.

15 The carbonate ions are formed from dissolved carbon dioxide, which reacts with water. Use this information to work out the equation for the formation of the carbonate ions.