

Chemistry 12  
Tutorial 3  
Separating Ions by Precipitation and  
Qualitative Analysis

In Tutorial 3 you will be shown:

1. How to use the Solubility Table to develop a scheme for *identification* of an unknown ion in a solution.
2. How to use the Solubility Table to outline a procedure to *separate* two ions from a solution through precipitation.

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### Qualitative Analysis

The term **Qualitative Analysis** means the use of *experimental procedures* to determine what elements or ions are present in a solution or substance.

In this unit, we will see how **differences in solubility** can help determine what an unknown ion is.

It's best to see how this is done using an example:

First, remember that on the Table:

"**Low Solubility**" means that a precipitate **will form**.

"**Soluble**" means that a precipitate **won't form**.

Let's say we have a solution which we know contains either  $\text{Ag}^+$  or  $\text{Ba}^{2+}$ , but we're not sure which. Now  $\text{Ba}^{2+}$  is known to be more poisonous, so we could get someone to drink it and see how fast they die. This, however is not advisable!

Another method would be to consult the Solubility Table and find something (a negative ion) which will form a precipitate with **one** of these ( $\text{Ag}^+$  or  $\text{Ba}^{2+}$ ) but **not** the other.

Let's locate both of these ions on the Table. The Positive Ions are all in the *center* column.

As you can see, both of these ions appear in several places

If both  $\text{Ag}^+$  &  $\text{Ba}^{2+}$  appear in the same group on the Table, that is not a good choice to use!

For example, look at the part of the Table with Sulphate ( $\text{SO}_4^{2-}$ ) on the left:

	All others	Soluble
Sulphate, $\text{SO}_4^{2-}$	<div style="display: flex; justify-content: space-around; align-items: center;"> <span><math>\text{Ag}^+</math>, <math>\text{Ca}^{2+}</math>, <math>\text{Sr}^{2+}</math>, <math>\text{Ba}^{2+}</math>, <math>\text{Pb}^{2+}</math></span> </div>	Low Solubility

( $\text{SO}_4^{2-}$ ) in the solu

What you must look for is something which forms a precipitate with **one** of them ( $\text{Ag}^+$  or  $\text{Ba}^{2+}$ ) but NOT with the other one!

Now look in the box *above* sulphate:

Ba<sup>2+</sup> is in the "All others" group.

Chloride, Cl <sup>-</sup> or Bromide, Br <sup>-</sup> or Iodide, I <sup>-</sup>	All others	Soluble
	Ag <sup>+</sup> , Pb <sup>2+</sup> , Cu <sup>+</sup>	Low Solubility

You can see, that in this case, chloride, bromide or iodide ions would form a precipitate with Ag<sup>+</sup> but NOT with Ba<sup>2+</sup>. The Ba<sup>2+</sup> is not in the bottom group, so therefore it must be one of "All others" It is soluble with Cl<sup>-</sup>.

So that means if you add some chloride (Cl<sup>-</sup>) ions to the solution and a precipitate forms, this tells you that the unknown ion **is Ag<sup>+</sup>** and **not Ba<sup>2+</sup>**. (NOTE: Remember we were told at the beginning that the ion is either Ag<sup>+</sup> or Ba<sup>2+</sup> and not something else! Analyzing a solution when you have no idea what ions are in it is a long complicated procedure involving many steps!)

Notice that chloride, bromide and iodide all act the same with these positive ions, so we could have added either chloride, bromide or iodide to get the same results.

Well, there is more than one way! See if you can do this question?

1. A solution is known to contain either Ag<sup>+</sup> or Ba<sup>2+</sup> ions. (One but not the other). Look on the Solubility Table in the box with "Sulphide" (S<sup>2-</sup>) on the left.

Sulphide, S <sup>2-</sup>	Alkali ions, H <sup>+</sup> , NH <sub>4</sub> <sup>+</sup> , Be <sup>2+</sup> , Mg <sup>2+</sup> , Ca <sup>2+</sup> , Sr <sup>2+</sup> , Ba <sup>2+</sup>	Soluble
	All others	Low Solubility

S<sup>2-</sup> ions are added to this solution and a precipitate DOES NOT FORM. Which ion, (Ag<sup>+</sup> or Ba<sup>2+</sup>) is present in the solution?

Answer \_\_\_\_\_

Often we are given more than just two possible ions in solution and have to come up with a procedure to identify them. Here's another example to follow through carefully, consulting your Solubility Table as you go along!

You are given three solutions in separate beakers. You are told that one solution contains Pb<sup>2+</sup> ions, one contains Ca<sup>2+</sup> ions and one contains Mg<sup>2+</sup> ions. Your job is to identify which is which using precipitation reactions.

Solution 1     
 Solution 2     
 Solution 3

One has  $Pb^{2+}$ , one has  $Ca^{2+}$  and one has  $Mg^{2+}$   
Which is which?

(COMMENT: There is often more than one method to solve these problems. You may think of another combination of things to add which might work just as well. I am just showing you one method that will work.)

The first job will be to find something which forms a precipitate (has "Low Solubility" with) only one of the three ions.

Have a look at the "Chloride" box on the Solubility Table:

Chloride, $Cl^-$ or Bromide, $Br^-$ or Iodide, $I^-$	All others	Soluble
	$Ag^+$ , $Pb^{2+}$ , $Cu^+$	Low Solubility

Notice that for our three ions ( $Pb^{2+}$ ,  $Ca^{2+}$ , and  $Mg^{2+}$ ), **only  $Pb^{2+}$**  forms a precipitate with chloride.

So what we could do is take a small sample from each solution (in little test tubes) and add chloride ( $Cl^-$ ) ions to each sample. Since  $Pb^{2+}$  is the only one of the ions which forms a precipitate with  $Cl^-$ , the sample with the precipitate (the one which turns cloudy) must be the one with the  $Pb^{2+}$  ions!

What we do at this point is set this solution aside (we have now identified it as the one with  $Pb^{2+}$ ).

Now, we look at *the other two solutions*. (one containing  $Ca^{2+}$  and one containing  $Mg^{2+}$ )

We look at our Solubility Table and try to find a negative ion which will form a precipitate with only one of these. ( $Ca^{2+}$  or  $Mg^{2+}$ )

Look at the "Sulphate" box:

Sulphate, $SO_4^{2-}$	All others	Soluble
	$Ag^+$ , $Ca^{2+}$ , $Sr^{2+}$ , $Ba^{2+}$ , $Pb^{2+}$	Low Solubility

We see that only  $Ca^{2+}$  (not  $Mg^{2+}$ ) forms a precipitate with sulphate. (*Low Solubility*)

So we take small samples of the remaining two solutions and add some sulphate ( $SO_4^{2-}$ ) ions to both.

The sample which forms the precipitate with sulphate must have the  $Ca^{2+}$  ions in it!

Now, by process of elimination, *the remaining solution* (the one which didn't precipitate with chloride or sulphate), must be the one with the  $Mg^{2+}$  ions.

So now we have identified which solution has  $Pb^{2+}$ , which has  $Ca^{2+}$ , and which has  $Mg^{2+}$ .

Note that it was best to take small samples of each one to do the experiments on. Once a precipitate forms in a beaker, it's hard to use the solution again for tests.

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A further complication:

In the example you just went through, it talked about "adding  $Cl^-$  (chloride) ions or adding  $SO_4^{2-}$  (sulphate) ions".

In reality, you can't just add those ions because *they don't exist by themselves* in a solution. Since all solutions are *neutral*, for every negative charge, there must be a positive charge.

So in order to "add  $Cl^-$  ions", you must add a solution which contains  $Cl^-$  ions.

*It is important that you always chose a soluble compound of the ion which you want to add in this type of activity.*

For example, if you wanted to add  $Cl^-$  ions to a solution, you must choose a soluble compound of  $Cl^-$ .

Furthermore, the positive ion should not form any unpredicted precipitates with anything already in the solution! (This would greatly complicate things. You don't need this. These are complicated enough as it is!)

In order to make sure the *positive ion* in the  $Cl^-$  solution doesn't do anything nasty, make sure it is an ion which is **soluble with everything!** Then you don't have to worry about it forming any unwanted precipitates.

Now, look on your Solubility Table.

All

Alkali ions:  $Li^+$ ,  $Na^+$ ,  $K^+$ ,  $Rb^+$ ,  $Cs^+$ ,  $Fr^+$

Soluble

Positive ions which are soluble with everything include the *Alkali Ions* ( $Li^+$ ,  $Na^+$ ,  $K^+$ ,  $Rb^+$ ,  $Cs^+$ ,  $Fr^+$ ), *Hydrogen* ( $H^+$ ) and *Ammonium* ( $NH_4^+$ ).

So the bottom line is, if a person wanted to add  $Cl^-$  ions to a solution, she/he could add an aqueous solution of  $NaCl$ ,  $KCl$ ,  $NH_4Cl$  etc. In these cases the  $Na^+$ ,  $K^+$ ,  $NH_4^+$  etc. would act as spectators and wouldn't interfere with anything.

Sodium compounds are usually quite common and easy to obtain. Also, you're probably used to using these. So, as a general rule:

If you need to use a *negative* ion, choose the sodium compound of it. Make sure the *formula is correct*. I promise you, you will lose marks for incorrect formulas!! (Use ion table!)

Let's say you wanted to add  $\text{SO}_4^{2-}$  (sulphate) ions to a solution. You could add an aqueous solution of **sodium sulphate** ( $\text{Na}_2\text{SO}_4$ ). This would supply the  $\text{SO}_4^{2-}$  ions and the  $\text{Na}^+$  ions present wouldn't interfere with anything. (They are spectators!)

In many questions, you are asked specifically what **compound** you would add - not just what ion! Make sure you always answer the question you are asked!

Let's say you need to add a positive ion to a solution in order to form a precipitate and test for something. Here's an example:

A solution is known to contain either  $\text{Cl}^-$  (chloride) ions or  $\text{SO}_4^{2-}$  (sulphate) ions. Name a **compound** which you could add to tell you which ion was present.

(Notice they said "compound" here, not just "ion"! Read the question again!)

Let's look at the Chloride and the Sulphate section on your Solubility Table:

Chloride, $\text{Cl}^-$ or Bromide, $\text{Br}^-$ or Iodide, $\text{I}^-$	All others	Soluble
	$\text{Ag}^+$ , $\text{Pb}^{2+}$ , $\text{Cu}^+$	Low Solubility
Sulphate, $\text{SO}_4^{2-}$	All others	Soluble
	$\text{Ag}^+$ , $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , $\text{Ba}^{2+}$ , $\text{Pb}^{2+}$	Low Solubility

Notice that  $\text{Ag}^+$  and  $\text{Pb}^{2+}$  ions form precipitates with both  $\text{Cl}^-$  and  $\text{SO}_4^{2-}$ !

So it wouldn't do any good to add  $\text{Ag}^+$  or  $\text{Pb}^{2+}$  ions!

But as you can see, **calcium** ions will form a precipitate with *sulphate* but **not** with *chloride*. So you could add calcium ions. (Strontium or barium would also work, but since calcium is more common, I'm choosing it. It doesn't really matter).

BUT, we can't just say to add calcium ions. The question asks us specifically for a **compound**. So we can add a compound of calcium.

But, guess what, it must be a compound in which the negative ion is soluble with everything.

Look at the Solubility Table. Negative ions which are soluble with everything are **nitrate** ( $\text{NO}_3^-$ ). So:

If you need to use a *positive* ion, choose the **nitrate** compound of it. **Make sure the formula is correct!** (Use ion table!)

In our example, we need to add calcium ( $\text{Ca}^{2+}$ ) ions, so the compound we could add is **calcium nitrate** ( $\text{Ca}(\text{NO}_3)_2$ ) (Make sure the formula is correct if you prefer full marks!)

This will supply the needed  $\text{Ca}^{2+}$  ions and the nitrate ( $\text{NO}_3^-$ ) ions are spectators and will not interfere with anything.

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2. A solution is known to contain either  $\text{Ba}^{2+}$  or  $\text{Mg}^{2+}$  ions.

Suggest a method by which these solutions could be analyzed to find out which ion is present. Be specific about any *compounds* that are added.

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3. A solution is known to contain one of these ions:  $\text{Mg}^{2+}$ ,  $\text{Ca}^{2+}$ ,  $\text{Sr}^{2+}$ , or  $\text{Be}^{2+}$ . Mixing samples of the solution with various reagents gives the following data:

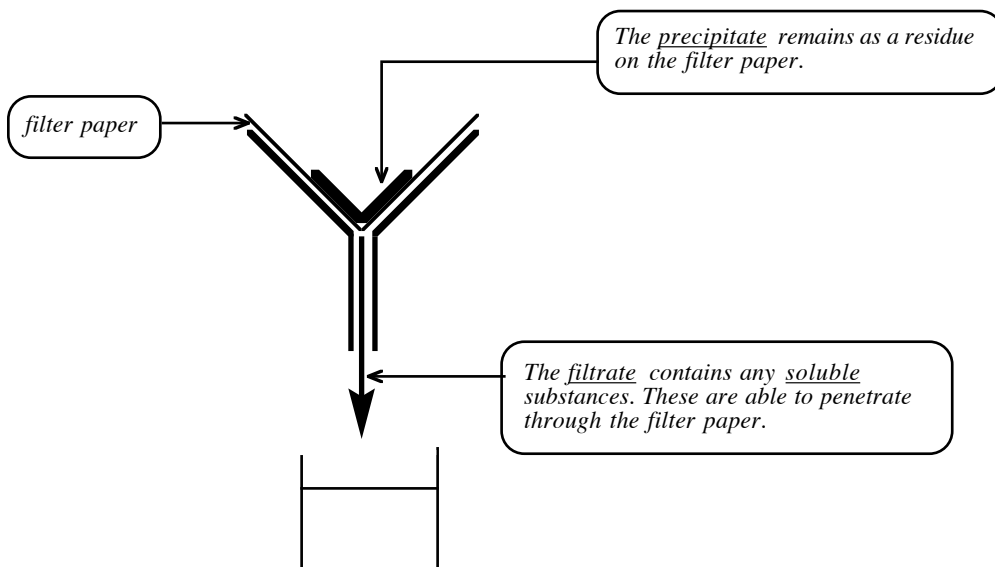
<b>Reagent</b>	$\text{Na}_2\text{S}$	$\text{Na}_2\text{SO}_4$	$\text{NaOH}$
<b>Result</b>	no ppt.	ppt.	no ppt.

From these data, which one of the four ions is present?

### Using Solubility to Separate Ions from Solutions

The second part of this Tutorial deals with *separating* ions from solutions which contain mixtures of ions using the Solubility Table.

One thing you have to realize in this type of problem is that once you form a *precipitate*, you can *separate* it from the rest of the solution by filtration:



The *precipitate* will be trapped in the filter paper and any soluble substances will still be dissolved in water and will go through the filter paper as the *filtrate*.

If there are more ions to precipitate, substances could be added to the filtrate and again precipitates could be filtered out.

Example:

You are given a solution containing *silver* ions ( $\text{Ag}^+$ ), *strontium* ions ( $\text{Sr}^{2+}$ ), and *magnesium* ions ( $\text{Mg}^{2+}$ ). You must separate them, one at a time from the solution using precipitation reactions. Show how you would do this. Write the Net-Ionic Equation for each precipitate formed. First of all, we must look at our Solubility Table and find something which precipitates only one of these three ions ( $\text{Ag}^+$ ), ( $\text{Sr}^{2+}$ ), or ( $\text{Mg}^{2+}$ ). Looking at the "Chloride Box"...

Chloride, $\text{Cl}^-$ or Bromide, $\text{Br}^-$ or Iodide, $\text{I}^-$	All others	Soluble
	$\text{Ag}^+$ , $\text{Pb}^{2+}$ , $\text{Cu}^+$	Low Solubility

We see that Chloride ions ( $\text{Cl}^-$ ) will precipitate *only*  $\text{Ag}^+$ , not  $\text{Sr}^{2+}$  or  $\text{Mg}^{2+}$ .

A good compound of Chloride to use would be sodium chloride ( $\text{NaCl}$ ). A 1 M solution is usually about the right concentration to use to get a precipitate, so here goes the first direction:

- 1. Add some 1 M  $\text{NaCl}$  solution (source of  $\text{Cl}^-$ ) to the container until no more precipitate forms.** (The  $\text{Cl}^-$  will precipitate the  $\text{Ag}^+$  from the solution.)



Filtration will keep the  $\text{AgCl}$  in the filter paper and the filtrate will not have any  $\text{Ag}^+$  ions left in it. (The  $\text{Cl}^-$  ions have bonded with all the  $\text{Ag}^+$  ions to form the precipitate,  $\text{AgCl}(\text{s})$ .) So the next direction is:

- 2. Filter the solution.  $\text{AgCl}(\text{s})$  will remain in the filter paper.**

The filtrate still has  $\text{Sr}^{2+}$  and  $\text{Mg}^{2+}$  ions in it, so now you have to find something which will precipitate just one of these. There are a couple of choices here. One choice is sulphate ( $\text{SO}_4^{2-}$ ):

Sulphate, $\text{SO}_4^{2-}$	All others	Soluble	The sulphate ( $\text{SO}_4^{2-}$ )
	$\text{Ag}^+$ , $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , $\text{Ba}^{2+}$ , $\text{Pb}^{2+}$	Low Solubility	

ions will precipitate the  $\text{Sr}^{2+}$  ions, but not the  $\text{Mg}^{2+}$  ions. (These are one of "All others" which would be soluble with sulphate.)

A good compound of Sulphate to use would be sodium sulphate ( $\text{Na}_2\text{SO}_4$ ). Remember to use the ion table to determine the correct formula!

So the next direction would read:

- 3. To the filtrate, add 1 M  $\text{Na}_2\text{SO}_4$  solution until a precipitate stops forming.**

The  $\text{SO}_4^{2-}$  from the  $\text{Na}_2\text{SO}_4$  will precipitate the Strontium ( $\text{Sr}^{2+}$ ) and form strontium sulphate. ( $\text{SrSO}_4$ ) - Check formula on green sheet!

The Net-Ionic equation would be:  $\text{Sr}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{SrSO}_4(\text{s})$

**Filtration** will remove the precipitate ( $\text{SrSO}_4(\text{s})$ ) from the solution:

So the next direction is:

**4. Filter the solution.  $\text{SrSO}_4(\text{s})$  will remain in the filter paper.**

The filtrate from this precipitation has only  $\text{Mg}^{2+}$  ions in it. (The other two have now been removed.)

To remove the  $\text{Mg}^{2+}$  ions, find something which forms a precipitate with  $\text{Mg}^{2+}$ .

(NOTE: Even if this ion forms a precipitate with  $\text{Ag}^+$  and/or  $\text{Sr}^{2+}$ , it doesn't matter because those two are **now gone**.) It's still better to choose something which forms a precipitate with  $\text{Mg}^{2+}$  ions only if that's possible.

Look at the "Hydroxide" box:

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	Alkali ions, $\text{H}^+$ , $\text{NH}_4^+$ , $\text{Sr}^{2+}$	Soluble
Hydroxide, $\text{OH}^-$	-----	
	All others	Low Solubility

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Hydroxide ( $\text{OH}^-$ ) forms a precipitate with  $\text{Mg}^{2+}$ , so by adding, let's say, sodium hydroxide, the  $\text{OH}^-$  will precipitate the  $\text{Mg}^{2+}$  and remove it from solution.

NOTE: You may notice that the  $\text{OH}^-$  also forms a precipitate with  $\text{Ag}^+$  ions. That doesn't matter in this case because **all the  $\text{Ag}^+$  ions have already been removed at this point**.

One thing you can see from this is that the order of adding things is important, so you really have to think it through!

So the next direction is:

**5. To the filtrate, add 1 M sodium hydroxide (NaOH) until the precipitate stops forming.**

The  $\text{OH}^-$  will precipitate the  $\text{Mg}^{2+}$ . We can now filter this to separate the Magnesium ions from the solution.

**The Net-Ionic Equation is:**  $\text{Mg}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Mg}(\text{OH})_2(\text{s})$

**6. Filter the solution to remove the  $\text{Mg}(\text{OH})_2(\text{s})$  precipitate.**

The original mixture, now has all three positive ions ( $\text{Ag}^+$ ), ( $\text{Sr}^{2+}$ ), & ( $\text{Mg}^{2+}$ ) removed by precipitation reactions.