Key

SCAN

Unit 3 – Part 4 – Qualitative Analysis:

Goals:

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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The term ***Qualitative Analysis*** means the use of *experimental procedures* to determine what elements or ions are present in a solution or substance. We will see how ***differences in solubility*** can help determine what an unknown ion is.

Find the "Solubility of some common ions" table in your data booklet.

"***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 Ag^+ or Ba^{2+}, but we're not sure which. We can consult the Solubility Table and find something (a negative ion) which will form a precipitate with ***one*** of these (Ag^+ or Ba^{2+}) but ***not*** the other.

Ba²⁺ is in the "All others" group.

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

So that means if you add some chloride (Cl^-) ions to the solution and a precipitate forms, this tells you that the unknown ion ***is Ag^+*** and ***not*** Ba^{2+}. 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.

There is more than one way! See if you can do this question?

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

Sulphide, S^{2-}	Alkali ions, H^+, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}	Soluble
	All others Ag.	Low Solubility

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

When Ba^{2+} is combined with S^{2-} it is soluble which means that it will not form a precipitate.

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

The first job will be to find something which forms a precipitate (has "Low Solubility" with) ***<u>only one</u>*** 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, ***only Pb^{2+}*** forms a precipitate with chloride. 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!

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 *<u>one of these</u>*. (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 <u>soluble</u> compound of the ion which you want to add in this type of activity.* *

Furthermore, the positive ion should not form any unpredicted precipitates with anything already in the solution! 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 ***<u>soluble with everything</u>***! Then you don't have to worry about it forming any unwanted precipitates.

Now, look on your Solubility Table.

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 ***<u>sodium</u>*** compound of it. Make sure the ***formula is correct***. (Use ion table!)

Example: A solution is known to contain either Cl^- (chloride) ions or SO_4^{2-} (sulphate) ions. Name a ***compound*** which you could add to tell you which ion was present.

Chloride, Cl^- or Bromide, Br^- or Iodide, I^-	All others	Soluble
	$\boxed{Ag^+, Pb^{2+}}$ Cu^+	Low Solubility
Sulphate, SO_4^{2-}	All others	Soluble
	$\boxed{Ag^+}$ Ca^{2+}, Sr^{2+}, Ba^{2+}, $\boxed{Pb^{2+}}$	Low Solubility

Notice that Ag^+ and Pb^{2+} ions form precipitates with <u>both</u> Cl^- and SO_4^{2-}! So it wouldn't do any good to add Ag^+ or Pb^{2+} ions! But as you can see, ***<u>calcium</u>*** ions will form a precipitate with *sulphate* but **<u>not</u>** with *chloride*. Look at the Solubility Table. Negative ions which are soluble with everything are ***nitrate*** (NO_3^-).

soluble with all:

$\boxed{Na^+ \quad NO_3^-}$

So, we can add Calcium nitrate ($Ca(NO_3)_2$)

2. A solution is known to contain either Ba^{2+} or 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.

Add $Na_2SO_4(aq)$ to the solution, if a precipitate is formed, the solution contains Ba^{2+}. ($BaSO_4$ has low solubility but $MgSO_4$ does not.)

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

Reagent	Na_2S	Na_2SO_4	NaOH
Result	no ppt.	ppt.	no ppt.

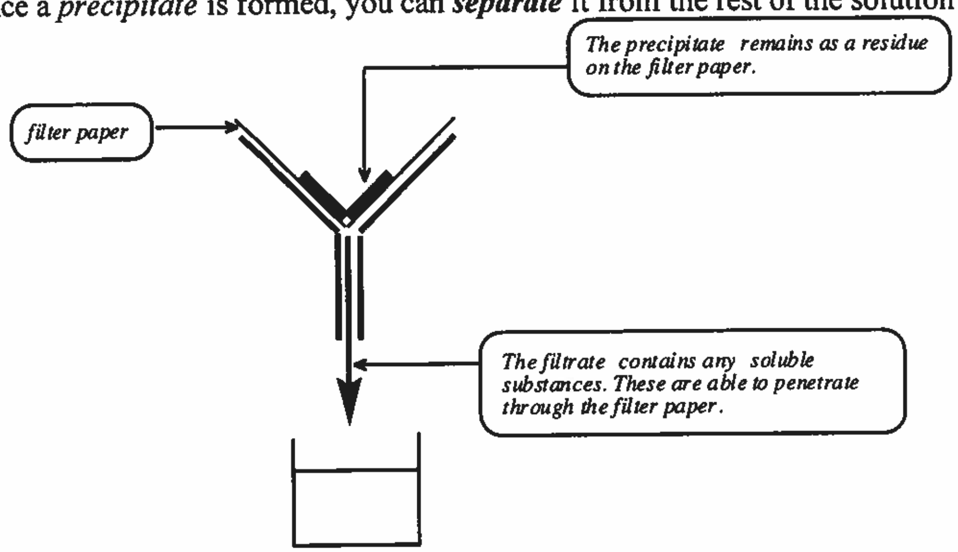
From these data, which one of the four ions is present?

→ if no ppt with S^{2-} then not Be^{2+}
→ if ppt with SO_4^{2-} then not Mg^{2+}
→ if no ppt with OH^- then not Ca^{2+}
→ Must be Sr^{2+}

Sr^{2+} is present in the solution

Using Solubility to Separate Ions from Solutions

Once a *precipitate* is formed, 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 (Ag^+), *strontium* ions (Sr^{2+}), and *magnesium* ions (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 (Ag^+), (Sr^{2+}), or (Mg^{2+}).

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

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

A good compound of Chloride to use would be sodium chloride (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 NaCl solution (*source of Cl^-*) to the container until no more precipitate forms. (*The Cl^- will precipitate the Ag^+ from the solution.*)

The Net-Ionic equation would be: $Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}$

2. Filter the solution. $AgCl_{(s)}$ will remain in the filter paper.

The filtrate still has Sr^{2+} and Mg^{2+} ions in it.

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

The sulphate (SO_4^{2-}) ions will precipitate the Sr^{2+} ions, but not the Mg^{2+} ions. A good compound of Sulphate to use would be sodium sulphate (Na_2SO_4).

3. To the filtrate, add 1 M Na_2SO_4 solution until a precipitate stops forming.
The SO_4^{2-} from the Na_2SO_4 will precipitate the Strontium (Sr^{2+}) and form strontium sulphate.

The Net-Ionic equation would be: $Sr^{2+}_{(aq)} + SO_4^{2-}_{(aq)} \rightarrow SrSO_{4(s)}$

4. Filter the solution. $SrSO_{4(s)}$ will remain in the filter paper

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

Hydroxide, OH^-	Alkali ions, H^+, NH_4^+, Sr^{2+}	Soluble
	All others	Low Solubility

Mg- is one of "All others"

Hydroxide (OH^-) forms a precipitate with Mg^{2+}, so by adding sodium hydroxide, the OH^- will precipitate the Mg^{2+} and remove it from solution.
*(<u>NOTE</u>: You may notice that the OH^- also forms a precipitate with Ag^+ ions. That doesn't matter in this case because **all the Ag^+ ions have already been removed at this point**.)*

5. To the filtrate, add 1 M sodium hydroxide (NaOH) until the precipitate stops forming.
The OH^- will precipitate the Mg^{2+}. We can now filter this to separate the Magnesium ions from the solution.

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

6. Filter the solution to remove the $Mg(OH)_{2(s)}$ precipitate.

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

<u>SELF TEST:</u>

1. You have 3 unlabelled test tubes containing I^-, Cu^{2+}, and Ca^{2+}. What procedures could you use to test these and find out which is which?

There are other possible answers!

* Add $NaI(aq)$, if a ppt forms, its Cu^{2+}
* Add $AgNO_3(aq)$, if a ppt forms, its I^-
* the remains test tube is Ca^{2+}

2. A solution contains both SO_4^{2-} and OH^-. Outline an experimental procedure to ***remove*** each ion ***individually*** from the solution and identify the reagents (***compounds***) used in this procedure. Include ***net-ionic equations*** for any precipitates formed.

Add $Sr(NO_3)_2(aq)$ and filter out the $SrSO_4(s)$.
Add $Ca(NO_3)_2(aq)$ and filter out the $Ca(OH)_2(s)$.

net ionic equations

$Sr^{2+}(aq) + SO_4^{2-}(aq) \rightarrow SrSO_4(s)$

$Ca^{2+}(aq) + 2OH^-(aq) \rightarrow Ca(OH)_2(s)$