



# 16

## Qualitative Analysis

### Introduction

The purpose of qualitative analysis is to determine what substances are present in detectable amounts in a sample. This experiment has two parts. In the first part, you will analyze an unknown solution for the presence of seven common ions. In the second part, you will test an unknown solid to determine which of two possible identities is correct.

### Part I. Spot Tests for Some Common Ions

A simple approach to the qualitative analysis of an unknown solution is to test for the presence of each possible ion by adding a reagent which will cause the ion, if it is in the sample, to react in a characteristic way. This method involves a series of “spot” tests, one for each ion, carried out on separate samples of the unknown solution. The difficulty with this way of doing qualitative analysis is that frequently, particularly in complex mixtures, one species may interfere with the analytical test for another. Although interferences are common, there are many ions which can be identified in mixtures by simple spot tests.

In this experiment we will use spot tests for the analysis of a mixture which may contain the following commonly encountered ions in solution:

$\text{CO}_3^{2-}$	carbonate
$\text{SO}_4^{2-}$	sulfate
$\text{PO}_4^{3-}$	phosphate
$\text{SCN}^-$	thiocyanate
$\text{Cl}^-$	chloride
$\text{C}_2\text{H}_3\text{O}_2^-$	acetate
$\text{NH}_4^+$	ammonium

The procedures we involve simple acid-base, precipitation, complex ion formation or oxidation-reduction reactions.

You will carry out each test three times. First, you will test 1 M solutions of each ion, so you can easily observe the expected results. Then you will test 0.1 M solutions of

each ion, which will require more careful observation. Finally, you will test an unknown solution, which will be a dilute mixture of 3 or 4 of the ions.

**SAFETY PRECAUTIONS:** Wear your SAFETY GOGGLES. If you spill any acid or base on your hands or clothing, rinse it off immediately with copious amounts of running water.

**WASTE DISPOSAL:** Pour all waste solutions into the INORGANIC WASTE containers in the fume hood.

**General Procedure:** Carry out the test for each of the ions as described below. Repeat each test using a solution made by diluting the anion solution 9:1 with distilled water; use your 10 mL graduated cylinder to make the dilution and make sure you mix well before taking the sample for analysis. In some of the tests, a boiling water bath will be needed, so set that up before proceeding. When performing a test, if no reaction is immediately apparent, stir the mixture with your stirring rod to mix the reagents. These tests can easily be used to detect the anions at concentrations of 0.02 M or greater, but in dilute solutions, careful observation may be required.

**Carbonate Ion.** Carbonate ion reacts with a strong acid to produce bubbles of carbon dioxide gas. Carbon dioxide is colorless and odorless.

*Procedure:*

Cautiously add 1 mL of 6 M HCl to 1 mL of 1 M Na<sub>2</sub>CO<sub>3</sub> in a small test tube. With concentrated solutions, bubbles of carbon dioxide gas are immediately evolved.

Repeat, using diluted Na<sub>2</sub>CO<sub>3</sub>. The effervescence will be much less obvious. Warming in the water bath, with stirring, will increase the rate of bubble formation.

*Questions:*

1. Write the net ionic equation for the reaction that occurs when the HCl solution is added to the Na<sub>2</sub>CO<sub>3</sub> solution.

**Sulfate Ion.** Sulfate ions react with barium ions to form insoluble barium sulfate, which is a white, finely divided precipitate. While two of the other anions in this experiment also form precipitates with barium, addition of HCl makes these compounds soluble. Barium sulfate is the only precipitate of barium that is not soluble in strong acid.

*Procedure:*

Add 1 mL of 6 M HCl to 1 mL of 0.5 M Na<sub>2</sub>SO<sub>4</sub>. Add a few drops of 1 M BaCl<sub>2</sub>.

Repeat, using diluted Na<sub>2</sub>SO<sub>4</sub>.

*Questions:*

2. Write the net ionic equation for the reaction that occurs when the BaCl<sub>2</sub> solution is added to the Na<sub>2</sub>SO<sub>4</sub> solution.

3. Which of the other anions would also form precipitates if the  $\text{BaCl}_2$  solution were added without the  $\text{HCl}$ ? Write net ionic equations to show how  $\text{HCl}$  makes these barium salts soluble.

**Phosphate Ion.** Phosphate ions react with ammonium molybdate to produce a characteristic yellow precipitate, ammonium phosphomolybdate  $(\text{NH}_4)_3\text{PO}_4 \cdot 12 \text{MoO}_3$ .

*Procedure:*

Add 1 mL 6 M  $\text{HNO}_3$  to 1 mL of 0.5 M  $\text{Na}_2\text{HPO}_4$ . Then add 1 mL of 0.5 M  $(\text{NH}_4)_2\text{MoO}_4$  and stir thoroughly. The ammonium phosphomolybdate precipitate may form slowly, particularly in dilute solutions; if a solid precipitate does not appear promptly, put the test tube in the boiling water bath for a few minutes.

Repeat, using diluted  $\text{Na}_2\text{HPO}_4$ .

**Thiocyanate Ion.** Thiocyanate ions react with ferric ions to form the complex ion  $\text{FeSCN}^{2+}$ , which has a characteristic deep red color. Two of the other anions in this experiment form precipitates with ferric ions, which would prevent the formation of the complex ion. However, addition of nitric acid makes these compounds soluble, so the  $\text{FeSCN}^{2+}$  complex ion can be observed.

*Procedure:*

Add 1 mL 6 M  $\text{HNO}_3$  to 1 mL of 0.5 M  $\text{KSCN}$  and stir. Add one or two drops of 0.1 M  $\text{Fe}(\text{NO}_3)_3$ .

Repeat, using diluted  $\text{KSCN}$ .

*Questions:*

4. Write the net ionic equation for the reaction that occurs when the  $\text{Fe}(\text{NO}_3)_3$  solution is added to the  $\text{KSCN}$  solution.
5. Which of the other anions would form precipitates if the  $\text{Fe}(\text{NO}_3)_3$  solution were added without the nitric acid?

**Chloride Ion.** Chloride ions react with silver ions to form insoluble white silver chloride. Unfortunately, many interferences are possible in tests involving silver ions. Thiocyanate ions, and two of the other anions in this experiment also form precipitates with dilute silver solutions. While addition of nitric acid makes two of these compounds soluble, both silver chloride and silver thiocyanate are insoluble in strong acid. If thiocyanate ions are present, they must be removed by oxidation before carrying out the test for chloride ions. After thiocyanate is removed, only silver chloride can precipitate under the conditions of this test.

*Procedure:*

Add 1 mL of 6 M  $\text{HNO}_3$  to 1 mL of 0.5 M  $\text{NaCl}$ . Add 2-3 drops of 0.1 M  $\text{AgNO}_3$ .

Repeat, using diluted  $\text{NaCl}$ .

If your sample contains  $\text{SCN}^-$  ion, put 1 mL of the solution in a small 30 or 50 mL beaker and add 1 mL of 6 M  $\text{HNO}_3$ . Boil the solution gently until the volume is decreased to one-half its original value; this will oxidize the thiocyanate and remove the interference. Then add another mL of 6 M  $\text{HNO}_3$  and a few drops of  $\text{AgNO}_3$  solution, which will form white  $\text{AgCl}$  in the presence of  $\text{Cl}^-$  ion as before.

*Questions:*

6. Write the net ionic equation for the reaction that occurs when the  $\text{AgNO}_3$  solution is added to the  $\text{NaCl}$  solution.
7. Which of the other anions (other than  $\text{SCN}^-$ ) would form precipitates if the  $\text{AgNO}_3$  solution were added without the nitric acid?
8. The purpose of adding only a few drops of dilute  $\text{AgNO}_3$  is to avoid the precipitation of another slightly soluble salt containing yet another one of the other anions. What is this other salt that might precipitate if a lot of concentrated  $\text{AgNO}_3$  were used? (You may need to look at a table of  $K_{\text{sp}}$  values.)

**Acetate Ion.** Acetate ion reacts with strong acids to form acetic acid, which has the characteristic odor of vinegar.

*Procedure:*

Add 1 mL 3M  $\text{H}_2\text{SO}_4$  to 1 mL of 1 M  $\text{NaC}_2\text{H}_3\text{O}_2$  and stir. Cautiously smell the mixture. The intensity of the vinegar odor is enhanced if the tube is warmed for 30 seconds in the boiling water bath.

Repeat, using diluted  $\text{NaC}_2\text{H}_3\text{O}_2$ . If the vinegar odor is undetectable, start again with a fresh sample of diluted  $\text{NaC}_2\text{H}_3\text{O}_2$ . Boil the sample carefully to near dryness before adding the  $\text{H}_2\text{SO}_4$ .

*Questions:*

9. Write the net ionic equation for the reaction that occurs when the  $\text{H}_2\text{SO}_4$  solution is added to the sodium acetate ( $\text{CH}_3\text{COONa}$  or  $\text{NaC}_2\text{H}_3\text{O}_2$ ) solution.

**Ammonium Ion.** Ammonium ion reacts with strong bases to form ammonia, which has a characteristic odor.

*Procedure:*

Add 1 mL of 6 M  $\text{NaOH}$  to 1 mL of 0.5 M  $\text{NH}_4\text{Cl}$ . Cautiously smell the mixture to determine if ammonia is present.

For a more sensitive test, pour the 1 mL of sample into a small beaker, moisten a piece of red litmus paper and put it on the bottom of a watch glass; cover the beaker with the watch glass and gently heat the liquid to the boiling point. Do not boil it and be careful that no liquid comes in contact with the litmus paper. The litmus paper will gradually turn blue if it is exposed to the vapors of evolving  $\text{NH}_3$ . Remove the watch glass and try to smell the ammonia.

Repeat, using diluted  $\text{NH}_4\text{Cl}$ .

*Questions:*

10. Write the net ionic equation for the reaction that occurs when the NaOH solution is added to the  $\text{NH}_4\text{Cl}$  solution.

**Unknown.** Obtain an unknown from your instructor. Record the unknown number. Analyze it by applying the spot tests to several 1 mL portions. Record the purpose and results of each test that you do on your unknown. The unknown will contain 3 or 4 of the ions on the list, so your test for a given ion may be affected by the presence of others. When you think you have properly analyzed your unknown, you may, if you wish, make a “known” which has the composition you found and test it to see if it has the properties of your unknown.

*Analysis:*

For each ion in this experiment, state whether it is or is not present in your unknown. Explain how each determination is supported by your experimental observations.

## **Part II. Identification of Unknown Solids**

In this experiment you will design and perform experiments to identify ionic solids. You will make use of your knowledge of general solubility rules, acid-base chemistry, solubility and complex ion equilibria, and even some electrochemistry. A variety of pure ionic solids will be available in numbered bottles. Each bottle will have two chemical formulas written on it. Your problem will be to determine which of the two chemical formulas on the bottle is correct.

Many of the solids can be identified on the basis of their anions, so that you can use the spot tests in Part I to find the correct formula. In other cases the compounds may both be insoluble or may both have the same anion, so that other approaches will be necessary. One very useful procedure is to consider the solubilities of the two compounds in different reagents or the solubilities of products which can be made from those substances.

Probably the most important solvent to consider is water. One of the compounds may be soluble in water and the other may not be. The appendix following this experiment tabulates the solubility properties of many ionic solids in water and other common laboratory reagents. This information will allow you to make many decisions regarding solubilities of ionic substances under different conditions. For example, if you wished to distinguish  $\text{MgSO}_4$  from  $\text{BaSO}_4$  (Example 1), you would find in the appendix that  $\text{MgSO}_4$  is soluble in water. In the table, you would also find that  $\text{BaSO}_4$  is insoluble in all common reagents. Thus to distinguish  $\text{MgSO}_4$  from  $\text{BaSO}_4$ , we need merely to add water.

Differences in solubility in solvents other than water can often be used to good advantage. Although most carbonates, phosphates, hydroxides, and oxides are insoluble in water, they are soluble in 6 M solutions of strong acids. This property would allow

you to distinguish  $\text{BaSO}_4$  from  $\text{BaCO}_3$  (Example 2) or  $\text{CuS}$  from  $\text{CuO}$  (Example 3). Another useful distinguishing solvent is 6 M  $\text{NH}_3$ , which dissolves many insoluble compounds by forming complex ions with the cation in the compound. Given a sample that might be  $\text{Mg}(\text{OH})_2$  or  $\text{Zn}(\text{OH})_2$  (Example 4) both water insoluble, and both soluble in strong acids, one could dissolve  $\text{Zn}(\text{OH})_2$  in 6 M  $\text{NH}_3$ .  $\text{Mg}^{2+}$ , however, does not form an ammonia complex ion, so  $\text{Mg}(\text{OH})_2$  would not dissolve in excess 6 M  $\text{NH}_3$ .

**Table 16.1. Example Procedures for Distinguishing Between Two Possible Solids**

Example 1: $\text{MgSO}_4$ (S according to appendix) or $\text{BaSO}_4$ (I according to appendix)		
Procedure	Expected result for $\text{MgSO}_4$	Expected result for $\text{BaSO}_4$
Mix a small amount of solid with water.	Would dissolve to give a colorless solution.	Would not dissolve.
Example 2: $\text{BaSO}_4$ (I according to appendix) or $\text{BaCO}_3$ (A according to appendix)		
Procedure	Expected result for $\text{BaSO}_4$	Expected result for $\text{BaCO}_3$
Mix a small amount of solid with 6 M HCl.	Would not dissolve.	Would dissolve with formation of bubbles to give a colorless solution. $\text{BaCO}_3 (s) + 2 \text{H}_3\text{O}^+ (aq) \rightarrow \text{Ba}^{2+} (aq) + \text{H}_2\text{O} (l) + \text{CO}_2 (g)$
Example 3: $\text{CuO}$ (A according to appendix) or $\text{CuS}$ (O according to appendix)		
Procedure	Expected result for $\text{CuO}$	Expected result for $\text{CuS}$
Add a small amount of solid to 6 M HCl. Stir well, while heating in a water bath.	Would dissolve to give a blue solution. $\text{CuO} (s) + 2 \text{H}_3\text{O}^+ (aq) \rightarrow \text{Cu}^{2+} (aq) + 3 \text{H}_2\text{O} (l)$	Would not dissolve.
Example 4: $\text{Mg}(\text{OH})_2$ (A according to appendix) or $\text{Zn}(\text{OH})_2$ (A, B, N according to appendix)		
Procedure	Expected result for $\text{Mg}(\text{OH})_2$	Expected result for $\text{Zn}(\text{OH})_2$
Add a small amount of solid to 6 M $\text{NH}_3$ . Stir very well.	Would not dissolve.	Would dissolve to give a colorless solution. $\text{Zn}(\text{OH})_2 (s) + 4 \text{NH}_3 (aq) \rightarrow \text{Zn}(\text{NH}_3)_4^{2+} (aq) + 2 \text{OH}^- (aq)$
Example 5: $\text{MgCl}_2$ (S according to appendix) or $\text{CdCl}_2$ (S according to appendix)		
Procedure	Expected result for $\text{MgCl}_2$	Expected result for $\text{CdCl}_2$
Add a small amount of solid to water. Add a few drops of $(\text{NH}_4)_2\text{S}$ in fume hood.	Would dissolve to give a colorless solution. No change.	Would dissolve to give a colorless solution. A precipitate would form. $\text{Cd}^{2+} (aq) + \text{S}^{2-} (aq) \rightarrow \text{CdS} (s)$
Example 6: $\text{CuS}$ (O according to appendix) or $\text{PbS}$ (O according to appendix)		
Procedure	Expected result for $\text{CuS}$	Expected result for $\text{PbS}$
Add a small amount of solid to 6 M $\text{HNO}_3$ . Boil the solution in the fume hood.	Would dissolve to give a blue solution and finely divided solid sulfur.	Would dissolve to give a colorless solution and finely divided solid sulfur.

Cool, filter out the sulfur, and add a few drops of Na <sub>2</sub> SO <sub>4</sub> .	$3 \text{ CuS (s)} + 2 \text{ NO}_3^- + 8 \text{ H}^+ \rightarrow 3 \text{ Cu}^{2+} + 3 \text{ S (s)} + 2 \text{ NO (g)} + 4 \text{ H}_2\text{O (l)}$ No change.	$3 \text{ PbS (s)} + 2 \text{ NO}_3^- + 8 \text{ H}^+ \rightarrow 3 \text{ Pb}^{2+} + 3 \text{ S (s)} + 2 \text{ NO (g)} + 4 \text{ H}_2\text{O (l)}$ A precipitate would form. $\text{Pb}^{2+} (\text{aq}) + \text{SO}_4^{2-} (\text{aq}) \rightarrow \text{PbSO}_4 (\text{s})$
Example 7: NaOH (S according to appendix) or NaHSO <sub>4</sub>		
Procedure	Expected result for NaOH	Expected result for NaHSO <sub>4</sub>
Add a small amount of solid to water. Test solution with litmus paper.	Would dissolve to give a colorless solution. Would turn red litmus blue; OH <sup>-</sup> is a strong base.	Would dissolve to give a colorless solution. Would turn blue litmus red. $\text{HSO}_4^- (\text{aq}) + \text{H}_2\text{O (l)} \rightarrow \text{SO}_4^{2-} (\text{aq}) + \text{H}_3\text{O}^+ (\text{aq})$
Example 8: NaCl or NaBr		
Procedure	Expected result for NaCl	Expected result for NaBr
Add a small amount of solid to water. Add a few mL of chlorine water (Cl <sub>2</sub> ).	Would dissolve to give a colorless solution. No change.	Would dissolve to give a colorless solution. Would turn pale yellow-orange. $2 \text{ Br}^- (\text{aq}) + \text{Cl}_2 (\text{aq}) \rightarrow \text{Br}_2 (\text{aq}) + 2 \text{ Cl}^- (\text{aq})$

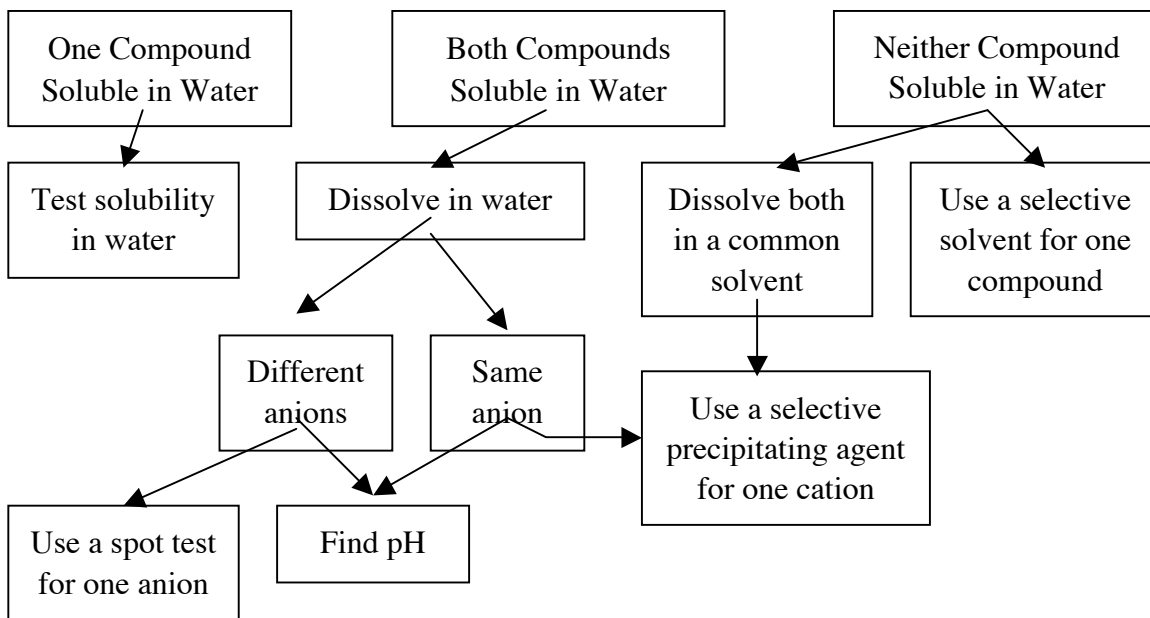
If both compounds are soluble in water or another common solvent and contain the same anion, then a precipitating reagent that reacts with one of the cations can be very helpful. Say, for example, you needed to distinguish between MgCl<sub>2</sub> and CdCl<sub>2</sub> (Example 5), both of which are water soluble. From the solubility tables, you could determine that MgS is water-soluble while CdS is not. Addition of 1 M (NH<sub>4</sub>)<sub>2</sub>S to the water solution of the unknown (in the fume hood!) should allow you to identify the substance present. In a more complicated case, with both compounds insoluble in water but soluble in a common solvent, it might be possible to find a selective precipitating reagent. If the unknown might be CuS or PbS (Example 6), considerable examination of solubility tables would reveal that, although both substances are very insoluble, they both dissolve in hot 6 M HNO<sub>3</sub>. Following treatment of the solid, which would put Cu<sup>2+</sup> or Pb<sup>2+</sup> into the solution, one could identify the ion by its color, since Cu<sup>2+</sup> is blue, or by precipitation on addition of 3 M H<sub>2</sub>SO<sub>4</sub>, which would precipitate PbSO<sub>4</sub> but not CuSO<sub>4</sub>.

Sometimes acidic or basic properties can be applied very easily to distinguish compounds. Anions like CO<sub>3</sub><sup>2-</sup> or OH<sup>-</sup> will always make their solutions basic. HSO<sub>4</sub><sup>-</sup> ion is highly acidic (Example 7). Color can sometimes be employed, as with the Cu<sup>2+</sup> in Example 6.

Although oxidation-reduction reactions are not often needed for distinguishing between compounds, they are sometimes very useful, as is the case with CuS and PbS (Example 6). Halide ions in solution can be treated with chlorine water, which will oxidize colorless Br<sup>-</sup> to pale yellow-orange Br<sub>2</sub> and colorless I<sup>-</sup> to dark yellow I<sub>2</sub> (Example 8). Nitric acid is a strong enough oxidizing agent to oxidize I<sup>-</sup>, but not Br<sup>-</sup>.



Although it is not feasible to present a procedure that will work for every possible pair of compounds, we have presented in the chart below a general outline which should be effective for dealing with most pairs. It essentially summarizes the material presented earlier in this discussion.



**SAFETY PRECAUTIONS:** Wear your SAFETY GOGGLES. If you spill any acid or base on your hands or clothing, rinse it off immediately with copious amounts of running water. Use extra caution in using the concentrated hydrochloric and nitric acid solutions. These are highly corrosive and should be used in the fume hood.

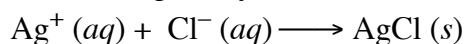
**WASTE DISPOSAL:** Pour all waste solutions into the INORGANIC WASTE containers in the fume hood.

## Procedure

In the fume hood, there will be a set of numbered bottles. Each bottle contains a pure solid compound, and is labeled with a number and the formulas of the two substances it may contain. Also available will be 6 M HCl, 6 M HNO<sub>3</sub>, 6 M NaOH, 6 M NH<sub>3</sub>,

12 M HCl, 12 M HNO<sub>3</sub>, chlorine water, and the reagents needed for the spot tests in Part 1. Design and carry out one- or two-step procedures using these reagents to identify the unknown solids. If at all possible, you should draw your conclusions from positive results. That is, your identifications should be based on reactions that did happen, rather than on reactions that did not happen.

There is one additional constraint on the procedures you design. You may use an identifying reaction only once. For example, if you use the reaction



to identify a substance, you may not use it again to distinguish between another pair of possibilities.

You should identify at least 8 unknowns. Work individually on this experiment. Your unknowns should be different from your normal lab partner's unknowns. You are welcome to consult with each other. For each unknown that you work on, you should record the unknown number, the two possible compounds, the procedures you follow, and your observations.

Waste disposal: All waste from this experiment should be poured into the inorganic waste bottles in the fume hood.

## Analysis

Give the unknown number and chemical formula of each compound you identified. Explain your reasoning for each identification, including net ionic equations for each reaction that actually occurred. (You may omit equations for dissolution of solids in water.)

## Pre-lab Question for Part 2

Devise one- or two- step procedures to determine which member of each pair is present in an unknown. For each procedure, write net ionic equations for any acid-base, precipitation, complex-forming, or redox reaction involved. (Use the examples in the earlier discussion as a model.)

- |  |   |
|--|---|
| (a) Na <sub>2</sub> CO <sub>3</sub> or NaCl  | (f) NaHSO <sub>4</sub> or Na <sub>2</sub> SO <sub>4</sub> |
| (b) NiCl <sub>2</sub> or BaCl <sub>2</sub>   | (g) MgCO <sub>3</sub> or BaSO <sub>4</sub>                |
| (c) KCl or NH <sub>4</sub> Cl  | (h) ZnCO <sub>3</sub> or CdCO <sub>3</sub>                |
| (d) NaOH or Mg(OH) <sub>2</sub>  | (i) AgCl or Ag <sub>2</sub> O                             |
| (e) KC <sub>2</sub> H <sub>3</sub> O <sub>2</sub> or Na <sub>2</sub> SO <sub>4</sub> | (j) ZnS or CuS  |

## Appendix: Solubility Properties of Ions and Solids

	Cl <sup>-</sup>	SO <sub>4</sub> <sup>2-</sup>	C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> <sup>-</sup>	CO <sub>3</sub> <sup>2-</sup> , PO <sub>4</sub> <sup>3-</sup>	OH <sup>-</sup> , O <sup>2-</sup>	S <sup>2-</sup>
Na <sup>+</sup> , K <sup>+</sup> , NH <sub>4</sub> <sup>+</sup>	S	S	S	S	S	S
Ba <sup>2+</sup>	S	I	S	A	S <sup>-</sup>	S
Ca <sup>2+</sup>	S	S <sup>-</sup>	S	A	S <sup>-</sup>	S
Mg <sup>2+</sup>	S	S	S	A	A	S
Fe <sup>3+</sup> (yellow)	S	S	S	A	A	A
Fe <sup>2+</sup> (pale aqua)	S	S	S	A	A	A
Cr <sup>3+</sup> (blue-violet)	S	S	S	A	A	A
Al <sup>3+</sup>	S	S	S	A, B	A, B	A, B
Ni <sup>2+</sup> (green)	S	S	S	A, N	A, N	A <sup>+</sup> , O <sup>+</sup>
Co <sup>2+</sup> (pink)	S	S	S	A	A	A <sup>+</sup> , O <sup>+</sup>
Zn <sup>2+</sup>	S	S	S	A, B, N	A, B, N	A
Mn <sup>2+</sup> (light pink)	S	S	S	A	A	A
Cu <sup>2+</sup> (blue)	S	S	S	A, N	A, N	O
Cd <sup>2+</sup>	S	S	S	A, N	A, N	A <sup>+</sup> , O
Ag <sup>+</sup>	A <sup>+</sup> , N	S <sup>-</sup> , N	S <sup>-</sup> , N	A, N	A, N	O
Pb <sup>2+</sup>	HW, B, A <sup>+</sup>	B	S	A, B	A, B	O

Note: Br<sup>-</sup> and I<sup>-</sup> have solubility properties similar to Cl<sup>-</sup>, but they have significantly different oxidation potentials.

**Key:** S soluble in water

S<sup>-</sup> slightly soluble in water

HW soluble in hot water

A soluble in acid (6 M HCl or other non-precipitating, non-oxidizing acid)

A<sup>+</sup> soluble in 12 M HCl

O soluble in hot 6 M HNO<sub>3</sub> (an oxidizing acid)

O<sup>+</sup> soluble in *aqua regia* (12 M HNO<sub>3</sub> and 12 M HCl mixed in a 1:3 ratio)

B soluble in 6 M NaOH

N soluble in 6 M NH<sub>3</sub>

I insoluble in any common reagent

## Interpreting the table of solubility properties

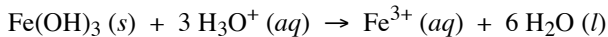
**S** The solubility in water is at least 0.1 M. For some compounds, shaking and stirring will be necessary. Compounds that are soluble in water are also soluble in acids, as long as the anion part of the acid does not precipitate with the cation part of the compound that you are trying to dissolve.

**S<sup>-</sup>** Some of the compound will dissolve in water, but its solubility is less than 0.1 M.

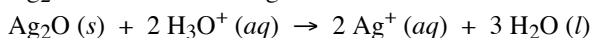
**HW** The compound will not dissolve in cold water, but its solubility in hot (near boiling) water is at least 0.1 M.

**A** The compound will not dissolve in pure water, but, in a 6 M solution of strong acid, its solubility is at least 0.1 M. An acid-base reaction is involved. Notice that most of the compounds in the table that are labeled with an A contain the conjugate bases of weak acids. When choosing the strong acid, make sure that the anion of the acid does not precipitate with the cation of the compound that you are trying to dissolve.

For example,  $\text{Fe}(\text{OH})_3$  is soluble in strong acid due to the acid-base reaction:



$\text{Ag}_2\text{O}$  is soluble in strong acid due to the acid-base reaction:

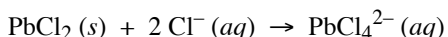


Why would you need to use  $\text{HNO}_3$  instead of  $\text{HCl}$  in the case of  $\text{Ag}_2\text{O}$ ?

When trying to react an oxide (rather than a hydroxide) with an acid, vigorous shaking and stirring and maybe even warming in a hot water bath will be necessary.

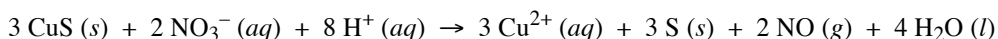
**A<sup>+</sup>** The compound will not dissolve in pure water or in 6 M acids, but, in a 12 M  $\text{HCl}$  solution, its solubility is at least 0.1 M. A complex ion formation reaction is involved.  $\text{Ag}^+$  and  $\text{Pb}^{2+}$  form stable complex ions with chloride ions. Vigorous shaking and stirring will be necessary. Be careful!

For example,  $\text{PbCl}_2$  is soluble in excess 12 M  $\text{HCl}$  due to formation of a complex ion:



**O** The compound will not dissolve in pure water, nor will it dissolve in acids via ordinary acid-base reactions, but it will react with hot 6 M  $\text{HNO}_3$  to form soluble ions. You will need to boil the solution. Use a boiling chip! An oxidation-reduction reaction is involved. Notice that most of the compounds in the table that are labeled with an O are sulfides. The sulfide ions ( $\text{S}^{2-}$ ) in these compounds will be oxidized to elemental sulfur (S).

For example, while  $\text{CuS}$  is insoluble in water and even in acids, it can be converted to soluble  $\text{Cu}(\text{NO}_3)_2$  by oxidation of its sulfide ion:

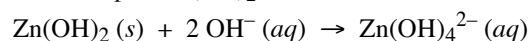


**O<sup>+</sup>** This is similar to O, except that the oxidizing reagents must be more concentrated for the oxidation-reduction reaction to go to completion. Hot 12 M  $\text{HNO}_3$  is needed in combination with 12 M  $\text{HCl}$ , which serves as an additional source of hydronium ions.

**B** The compound will not dissolve in pure water, but, in a 6 M solution of  $\text{NaOH}$ , its solubility is at least 0.1 M. A complex ion formation reaction is involved.  $\text{Al}^{3+}$ ,

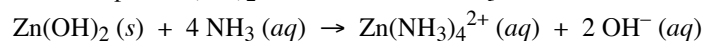
$\text{Zn}^{2+}$ , and  $\text{Pb}^{2+}$  form stable complex ions with hydroxide ions ( $\text{OH}^-$ ). Vigorous shaking and stirring will be necessary.

For example,  $\text{Zn}(\text{OH})_2$  is soluble in  $\text{NaOH}$  due to formation of a complex ion:



**N** The compound will not dissolve in pure water, but, in a 6 M solution of  $\text{NH}_3$ , its solubility is at least 0.1 M. A complex ion formation reaction is involved.  $\text{Ni}^{2+}$ ,  $\text{Zn}^{2+}$ ,  $\text{Cu}^{2+}$ ,  $\text{Cd}^{2+}$ , and  $\text{Ag}^+$  form stable complex ions with ammonia ( $\text{NH}_3$ ). Vigorous shaking and stirring will be necessary.

For example,  $\text{Zn}(\text{OH})_2$  is also soluble in  $\text{NH}_3$  due to the formation of another complex ion:



**I** The compound will not dissolve or react in any reagent available in the laboratory.

Note: The table can be used to predict the solubilities of *dry* solids in the various reagents. If you want to test a solid that has precipitated from a solution, you must first centrifuge the mixture and pour off the liquid.