Resolve to be thyself: and know,
that he Who finds himself, loses
his misery. --Matthew Arnold

Precipitation Reactions

Reactions occurring in solution may produce substances which are *insoluble* in the solution and thus eventually settle to the bottom or "precipitate" out. This is of interest to chemists because so much of what we do occurs in some type of solution, generally in water.

Typically **precipitation reactions** occur in *aqueous* solution between ions (i.e., the reactants are *electrolytes*). Often only one pair of ions actually react or precipitate while the other pair remain in solution unchanged. These ions which remain unchanged during the reaction are called *spectators* and can be omitted from the final balanced equation written in *net-ionic* form. As you will see...

Although solubility is often related to the position of an element in the periodic table, there are enough exceptions to make for an interesting (or annoying, depending on your point of view) set of rules that describe which combinations are insoluble. One object of this experiment is to arrive at a partial set of rules that describes which combinations of ions are insoluble in water solution.

Preparing to experiment

Attached to this sheet you will find a grid in which to fill in your observations of what happens when you mix pairs of solutions. A similar paper is waiting for you in the lab laminated in plastic. All of the reactions can be performed with drop quantities of each solution on top of the plastic sheet.

After you have recorded your results, carefully lift the plastic sheet and tip the drops into the sink. Rinse the plastic gently and carefully with water and dry with paper towels. Try not to crinkle it. Waste not, want not.

The grid is provided for your use as an observation table.

Technique

What may at first seem obvious is often not [write that down somewhere!]. To complete this experiment in a single period you need to have some kind of organized approach to putting so many drops onto the sheet--especially since you will probably be sharing a set of solutions with another student. Your instructors have given this a GREAT deal of thought and we recommend the following:

a. Pick up a solution and place drops of it in *each* square where it will be used. Remember that simply putting drops in a horizontal row or vertical column will NOT generally accomplish this since most substances are in *both* rows and columns. For example, if you look at the grid you will see that $Al(NO_3)_3$ (about half-way down the left side) turns up when you reach the end of the row. Most substances "turn a corner" like this.

b. Mark off the solutions you have used with a small check mark on your data sheet. Most solutions look alike.

c. Be careful not to touch the second solution dropper to the first drop of solution that has already been placed (why?)

d. Do not add Na_2S until **after you have recorded all the results for other mixtures**. It has an unpleasant smell and will also change the appearance of all your mixtures!

Analysis

These questions should be answered in your laboratory notebook following your observations.

1. So much data and so little time.... If you inspect your observation grid carefully, you should see that there is one compound on the grid which did not react at all. This was NH_4NO_3. Since **no** combination of ions reacted with either NH_4^+ or NO_3^-, we might say:

1) all compounds containing NH_4^+ are soluble
2) all compounds containing NO_3^- are soluble

These two rules should make it possible for you to figure out the rest. To help you do so, fill in the grid below. This is done by looking for compounds which contain the anion listed in the left column and placing a check mark in the column under a cation which *when added* caused a precipitate to form [note that the nitrate and ammonium sections have been "grayed-out"--no reactions occur there]. The Chloride ion is done for you as an example.

	Hg_2^{2+}	Ba^{2+}	Na^+	Ca^{2+}	K^+	Mg^{2+}	Cu^{2+}	NH_4^+	Al^{3+}	Ag^+	Pb^{2+}	Fe^{3+}
NO_3^-												
PO_4^{3-}												
S^{2-}												
CO_3^{2-}												
Br^-												
SO_4^{2-}												
Cl^-	✓									✓	✓	
OH^-												

2. Now select each anion represented in the experiment and write a rule for its solubility with the various cations you used. Write the rule in the most economical way. In other words, if there are more ions which are soluble with the anion, list those with which it is insoluble. Below is a reasoned example to get you started:

The compound KCl contains Cl^- ion and forms precipitates with only three compounds: $AgNO_3$, $Pb(NO_3)_2$ and $Hg_2(NO_3)_2$.

In each case one of the possible products is KNO_3. But we already know that all NO_3^- compounds are soluble. So it must be the chloride compounds that are precipitates. Thus:

3) Cl^- is insoluble with Ag^+, Pb^{2+} and Hg_2^{2+}

Continue in this way to write rules for the remaining **anions** (it may be helpful to proceed in a similar fashion, tackling those that show the *fewest* reactions first before trying the more complicated cases)

3. Many of the solutions you used in the experiment contain the anion **nitrate** (NO_3^-) and many also contain cations from **Group I** (e.g., Na^+, K^+). When your instructor prepares water solutions for experiments, why is it a good idea to choose compounds containing one of these ions?

4. Write *molecular* balanced equations for any 10 of the reactions which resulted in precipitates.

5. Write *net-ionic* balanced equations for the reactions you wrote in #4.

	$Ba(NO_3)_2$	$NaOH$	$Ca(NO_3)_2$	KCl	Na_2SO_4	$Mg(NO_3)_2$	$Cu(NO_3)_2$	$NaBr$	NH_4NO_3	$Al(NO_3)_3$	$AgNO_3$	$Pb(NO_3)_2$	Na_2CO_3	Na_2S	Na_3PO_4	$Fe(NO_3)_3$
$Hg_2(NO_3)_2$																
$Fe(NO_3)_3$																
Na_3PO_4																
Na_2S																
Na_2CO_3																
$Pb(NO_3)_2$																
$AgNO_3$																
$Al(NO_3)_3$																
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