# ­­Identifying Precipitates Name : \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Per: \_\_

**Purpose:** ­­

1. To identify ionic compounds that are not soluble in water
2. To predict the products of a double replacement reaction

**Background Information:** Some ionic solids are much more soluble in water than others. This means that some ionic solids will easily dissolve in water where as some will not. There are a number of ionic solids that hardly dissolve at all in water and we call them “insoluble”. When an ionic solid is dissolved in water, the ions contained in them are dissociated from each other and are surrounded by water. We use the word “aqueous” to describe this. Because they are aqueous they are free to interact with other oppositely charged ions if two solutions are mixed together. Sometimes a new combination leads to the formation of a compound that is not very soluble at all. The ions will come out of solution and form a solid (precipitate). This formation of a precipitate can be observed by the solution losing its clarity, and becoming cloudy or milky in appearance.

**Procedure:**

1. On the transparency provided, put two drops of the appropriate solution in each square.
2. Do not touch the dropper bottle to the transparency. To avoid contamination, **release each drop about an inch above the transparency.**
3. Allow each reaction to occur. If the products of the solution are not easily observed, add the two reactants again.
4. Observe each reaction. Write “**ppt**” (shorthand for precipitate) in each square where the products become cloudy. Look closely, it may be only a slight milkiness.

You will notice from looking at the grids that each combination of two solutions will be done twice. This is to give you an extra opportunity to pick up the faint precipitates. If the two results contradict each other, add the two reactants again. ***If no precipitate forms, leave the square blank.***

1. Since this is a micro-scale experiment, the amounts of solute are so small that you may dispose of them by rinsing off the overhead down the sink.
2. Dry transparency and return it and reagents to appropriate containers.

**Data Tables:**

**SET I**

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
|  | **AgNO3** | **Pb(NO3)2** | FeCl3 | **MgSO4** | **CuSO4** | **NH4Cl** |
| **NaNO3** |  |  |  |  |  |  |
| NaOH |  |  |  |  |  |  |
| **Na2CO3** |  |  |  |  |  |  |
| **Na2SO4** |  |  |  |  |  |  |

**Post Lab Questions:**

1. In each reaction when a precipitate formed write the equation with the two new products.

Example: NaOH + MgCl2 🡪 NaCl + Mg(OH)2

1. Then, Balance the reaction. Of the two products, one will be a precipitate and the other will be aqueous. Label which product is which with (s) or (aq). Using the previous example, we know that NaCl (table salt) dissolves in water, therefore, the Mg(OH)2 **must** be the precipitate. Decide which of your two products will be the **ppt** and which will be the solution, using your solubility table.

Example: 2NaOH(aq) + MgCl2(aq) 🡪 2NaCl ­(aq) + Mg(OH)2 (s)

**For all reactions that form a precipitate, write the balanced reactions with states of matter in the space below OR on a separate sheet of paper. *Circle the precipitate*.**

Example:

1. 2NaOH(aq) + MgCl2(aq) 🡪 2NaCl ­(aq) + Mg(OH)2 (s)