Predicting Aqueous Precipitation Reactions

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In this activity we will learn to be able to predict if a reaction occurs and a precipitate forms when two aqueous solutions are mixed, and if so, what that precipitate is. A salt that dissolves in water is called a soluble salt and forms an aqueous solution (homogenous mixture). One that does not dissolve is called an insoluble salt and forms a heterogenous mixture. A precipitation reaction occurs when aqueous solutions (two soluble salts) mix, undergo a double displacement reaction and an insoluble salt forms (solid precipitate). There are three ways of writing a precipitation reaction, as a complete (molecular) equation, total ionic equation or net ionic equation. You need to review your text on this subject (see learning guide). Note, no reaction occurs if all ions are spectator ions (you simply mixed two soluble salts and nothing happened).

Learning Objectives:

1. Predict products for single and double displacement reactions
2. Use solubility rules to predict if a salt is soluble (forms an aqueous solution) or insoluble (forms a solid precipitate)
3. Write general equations for single and double displacement reactions (include phases in parenthesis)
4. Write total ionic equations for single and double displacement reactions (treat aqueous ionic compounds as dissociated ions)
5. Write net ionic equations for single and double displacement reactions (ignore spectator ions, which are ions which exist on both sides of the equation)

Part 1. Determine what the products will be for the following reactions. Initially we will ignore phases, but by the time we finish this exercise you will need to be able to identify phases as you write the equation. This process requires that you identify the ions in the reactants and for double displacement reactions swap partners and use the principle of charge neutrality to determine what the product are. For single replacement reaction you need to figure which species (anion or cation) is gaining or losing charge and the other species is a spectator ion. Once that is done you need to balance the equations

Students frequently swap ions without thinking about the product formula and this leads to mistakes. It is suggested you follow these steps:

1. Identify reactant ions and their charges
2. Determine Product Formula based on principle of charge neutrality
3. Balance Equation

Double Displacement Reactions
<u>(In Class)</u>

a. ___$Al(NO_3)_3(aq)$ + ___$Na_2(CO_3)$ (aq) -->

b. ___$Pb(ClO_4)_2(aq)$ + ___$Na_2SO_4(aq)$ -->

(at home)

c. ___$(NH_4)_3PO_4$(aq) + ___NaOH(aq) -->

d. ___$(NH_4)_3PO_3$(aq) + ___$CaSO_4$(s) -->

e. ___$(NH_4)_2SO_4$(aq) + ___$CaCl_2$(aq) -->

f. ___$AgCH_3CO_2$ (aq) + ___ KCl(aq) -->

g. ___$Pb(NO_3)_2$(aq) + Na_2SO_4(aq) -->

h. ___NaOH + $CaCl_2$(aq) -->

i. ___BaI_2 + H_2SO_4(aq) -->

Single Displacement reactions
(in class)

j. Mg(s) + $FeCl_3$(aq) -->

k. $AgNO_3$(aq) + Ca(s) -->

(take home)

l. Na(s) + HCl(aq) -->

m. Na(s) + H_2O(l) -->

n. Zn(s) + AgCl(aq) --->

o. Ag(s) + $ZnCl_2$(aq) --->

Part 2: Predicting Solubility: You use the solubility rules to determine if an ionic compound is soluble or insoluble. You need to memorize these.

Solubility Rules

I. Soluble
- **a. Group 1A & Ammonium (the only cations in this list)**
- **b. NO_3^-, ClO_4^-, ClO_3^-, $CH_3CO_2^-$**

II. Usually Soluble
- **a. Cl^-, Br^-, I^-, (Except those with Ag^+, Hg_2^{+2}, & Pb^{+2})**
- **b. SO_4^{-2} (Except those with Hg^{+2}, Ca^{+2}, Sr^{+2}, Ba^{+2} & Pb^{+2})**

III Insoluble (Except with cations from I.A)
- **a. OH^- (Except those with Ca^{+2}, Sr^{+2}& Ba^{+2})**
- **b Everything else (this is not true, but will work in this class)**

Use the solubility rules to determine if the following compounds are soluble or insoluble. Indicate answer by writing formula followed by (aq) for soluble and (s) for insoluble. (You may also want to write the names of the species)

(in class)

1. $Cu(NO_3)_2$ 2. $Pb(NO_3)_4$

3. $PbCl_2$ 4. AgI

5. $PbSO_4$ 6.NaBr

(at home)

7. calcium carbonate 8. calcium nitrate

9. iron (III) chloride 10. chromium (VI) acetate

11. Barium sulfate 12. barium chloride

Part 3.

<u>Writing the Complete Equation</u> (often called molecular equation). Write the answers from Part 1 but now use the solubility rules to include the phases of each species in parenthesis after the formula. Remember, (aq) = aqueous, (s) = solid, (g) = gas and (l) = liquid. Be sure eq. is balanced

(In Class)

a. ___$Al(NO_3)_3$(aq) + ___$Na_2(CO_3)$(aq) -->

b. ___$Pb(ClO_4)_2$(aq) + ___Na_2SO_4(aq) -->

(at home)

c. ___$(NH_4)_3PO_4$(aq) + ___NaOH(aq) -->

d. ___$(NH_4)_3PO_3$(aq) + ___$CaSO_4$(s) -->

e. ___$(NH_4)_2SO_4$(aq) + ___$CaCl_2$(aq) -->

f. ___$AgCH_3CO_2$ (aq) + ___ KCl(aq) -->

g. ___$Pb(NO_3)_2$(aq) + Na_2SO_4(aq) -->

h. ___NaOH + $CaCl_2$(aq) -->

i. ___BaI_2 + H_2SO_4(aq) -->

Single Replacement Reactions

j. Mg(s) + $FeCl_3$(aq) -->

k. $AgNO_3$(aq) + Ca(s) -->

l. Na(s) + HCl(aq) -->

m. Na(s) + H_2O(l) -->

Part 4. **Complete Ionic Equation.** Under each equation write all ionic aqueous species as separate ions. For example, write NaCl(aq) as Na^+(aq) and Cl^-(aq), but write AgCl(s) as AgCl(s).

<u>(In Class): Double Displacement Reactions</u>

a. ___$Al(NO_3)_3$(aq) + ___$Na_2(CO_3)$(aq) -->

b. ___$Pb(ClO_4)_2$(aq) + ___Na_2SO_4(aq) -->

<u>(at home)</u>

c. ___$(NH_4)_3PO_4$(aq) + ___NaOH(aq) -->

d. ___$(NH_4)_3PO_3$(aq) + ___$CaSO_4$(s) -->

e. ___$(NH_4)_2SO_4$(aq) + ___$CaCl_2$(aq) -->

f. ___$AgCH_3CO_2$ (aq) + ___ KCl(aq) -->

g. ___$Pb(NO_3)_2$(aq) + Na_2SO_4(aq) -->

h. ___NaOH + $CaCl_2$(aq) -->

i. ___BaI_2 + H_2SO_4(aq) -->

Single Replacement Reactions

j. Mg(s) + $FeCl_3$(aq) -->

k. $AgNO_3$(aq) + Ca(s) -->

l. Na(s) + HCl(aq) -->

m. Na(s) + H_2O(l) -->

Part 5: Net Ionic Equation. Under each equation write the net ionic equation. You do this by cancelling out spectator ions from the total ionic equation. Note, the stoichimetric coefficent of the net ionic equation may be different than from the molecular or total ionic. **(In Class)**

a. ___$Al(NO_3)_3(aq)$ + ___$Na_2(CO_3)(aq)$ -->

b. ___$Pb(ClO_4)_2(aq)$ + ___$Na_2SO_4(aq)$ -->

(at home)

c. ___$(NH_4)_3PO_4(aq)$ + ___$NaOH(aq)$ -->

d. ___$(NH_4)_3PO_3(aq)$ + ___$CaSO_4(s)$ -->

e. ___$(NH_4)_2SO_4(aq)$ + ___$CaCl_2(aq)$ -->

f. ___$AgCH_3CO_2$ (aq) + ___ $KCl(aq)$ -->

g. ___$Pb(NO_3)_2(aq)$ + $Na_2SO_4(aq)$ -->

h. ___$NaOH$ + $CaCl_2(aq)$ -->

i. ___BaI_2 + $H_2SO_4(aq)$ -->

Single Replacement Reactions:

j. $Mg(s)$ + $FeCl_3(aq)$ -->

k. $AgNO_3(aq)$ + $Ca(s)$ -->

l. $Na(s)$ + $HCl(aq)$ -->

m. $Na(s)$ + $H_2O(l)$ -->

Part 6: Mixing Multiple Solutions:

What are the precipitates formed if multiple solutions mix?

1. Consider 3 aqueous salt solutions each with different cations and anions. For example, consider NaCl, $AgNO_3$ and $(NH_4)_2CO_3$. There are 3 cations and 3 anions which results in 3^2 or 9 combinations (of which 3 are the reactants). This seems like a very complicated problem (4 salts would result in 16 potential combinations and 5 salts would result in 25) and so we need to develop a technique to see the problem. This can be done through a matrix, where the rows represent the cations, the columns the anions and the cells the potential combinations:

	Cl^-	NO_3^-	CO_3^{-2}
Na^+			
Ag^+			
NH_4^+			

Now think about it for a minute. By writing the reactants out this way the diagonal represents your reactants and anything in the first or third row must be soluble (Rules 1A), so you can strike them out. Likewise anything in the second column must be soluble, so you have reduced this problem to two questions. Are silver chloride and silver carbonate precipitates or do they form aqueous solutions?

	Cl^-	NO_3^-	CO_3^{-2}
Na^+			
Ag^+			
NH_4^+			

Note, all sodium, nitrate and ammonium salts are soluble and so we can ignore them in identifying potential precipitates.

In Class:

Identify if any precipitates in the following solutions are mixed:

1. $Pb(ClO_4)_2$(aq), K_2SO_4(aq), $AgCH_3CO_2$ (aq) + KCl(aq)

2. $Ba(ClO_3)_2$(aq), Li_2SO_4(aq), $NH_4CH_3CO_2$ (aq) + $BaCl_2$(aq)