COLLEGE OF THE CANYONS

Activity 151-7
Writing Complete Ionic and Net Ionic Equations

Directions: This worksheet is focused on writing both complete ionic and net ionic equations. These types of equations are more commonly used for Precipitation and Acid/Base Reactions. Part A introduces writing compounds as they exist in aqueous solution. Part B discusses writing complete ionic equations by writing products and reactants as they exist in aqueous solution. Part C discusses writing balanced net ionic equations by removing spectator ions from the complete ionic equation. The worksheet is accompanied by instructional videos. See http://www.canyons.edu/Departments/CHEM/GLA for additional materials.

Part A – Writing compounds as they exist in aqueous solution

In both the complete ionic and net ionic equations, each compound is written as it exists in solution. There are some rules to follow:

1. If the ionic compound is soluble, (*aq*), it should be split up into its ions. Keep in mind that polyatomic ions are one unit and should remain as one unit with an overall charge (not split up). Refer to the "Chemical Formulas for Ionic Compounds" activity to help you determine what the charge on the transition metals will be.
2. If the ionic compound is insoluble, (*s*), it stays associated in solution, so it should not be split up into its ions. Refer to the "Predicting Products" activity to learn how to determine whether an ionic compound is soluble or insoluble.
3. Liquid water, H_2O (*l*), should also never be split up into its ions.

Practice:

$KCl(aq)$	$K^+(aq) + Cl^-(aq)$	$PbCl_2(s)$	$PbCl_2(s)$
$Pb(C_2H_3O_2)_2(aq)$	$Pb^{2+}(aq) + 2C_2H_3O_2^-(aq)$	$H_2O(l)$	$H_2O(l)$

Example #1:

a) $FeBr_3(aq)$	____________	f) $CoCO_3(s)$	____________
b) $NaBr(aq)$	____________	g) $K_2S(aq)$	____________
c) $AgCl(s)$	____________	h) $Pb(CO_3)_2(s)$	____________
d) $CoCl_2(aq)$	____________	i) $Li_2SO_4(aq)$	____________
e) $H_2O(l)$	____________	j) $SrCl_2(aq)$	____________

Part B – Writing complete ionic equations

Complete ionic equations are a type of equation that show how the species exist in solution. Just as in Part A, if the species is soluble, (*aq*), it will be split up into its ions and if it is insoluble, (*s*), it will remain as one unit. When these compounds are written in a balanced chemical equation, they may have a coefficient associated with them. In order to keep the complete ionic equation balanced, the coefficients also need to be included. If the soluble, (*aq*), ionic compound contains a coefficient, then the coefficient must be distributed to both the cation and that anion that are split up.

Practice:

$2\ KCl(aq) + Pb(C_2H_3O_2)_2(aq) \rightarrow 2\ KC_2H_3O_2(aq) + PbCl_2(s)$

Complete ionic:

$\mathbf{2\ K^+(aq) + 2\ Cl^-(aq) + Pb^{2+}(aq) + 2\ C_2H_3O_2^-(aq) \rightarrow 2\ K^+(aq) + 2\ C_2H_3O_2^-(aq) + PbCl_2(s)}$

Example #2: $3\ NaOH(aq) + FeBr_3(aq) \rightarrow Fe(OH)_3(s) + 3\ NaBr(aq)$

Example #3: $H_2SO_4(aq) + 2\ NaOH(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$

Part C – Writing net ionic equations

Net ionic equations show only the species that actually reacted in solution. This done by removing spectator ions from the complete ionic equation. A spectator ion is any species that is the same on both the reactants side and the products side. If all your products are aqueous, meaning all you have are spectator ions, then there is no reaction.

Practice (from Part B):

$2\ KCl(aq)\ +\ Pb(C_2H_3O_2)_2(aq)\ \rightarrow\ 2\ KC_2H_3O_2(aq)\ +\ PbCl_2(s)$

Complete ionic (from Part B):

$2\ K^+(aq)\ +\ 2\ Cl^-(aq)\ +\ Pb^{2+}(aq)\ +\ 2\ C_2H_3O_2^-(aq)\ \rightarrow\ 2\ K^+(aq)\ +\ 2\ C_2H_3O_2^-(aq)\ +\ PbCl_2(s)$

Net ionic:

$\cancel{2\ K^+(aq)}\ +\ 2\ Cl^-(aq)\ +\ Pb^{2+}(aq)\ +\ 2\ \cancel{C_2H_3O_2^-}(aq)\ \rightarrow\ \cancel{2\ K^+(aq)}\ +\ 2\ \cancel{C_2H_3O_2^-}(aq)\ +\ PbCl_2(s)$

$Pb^{2+}(aq)\ +\ 2\ Cl^-(aq)\ \rightarrow\ PbCl_2(s)$

Practice:

$BaCl_2(aq)\ +\ 2\ AgNO_3(aq)\ \rightarrow\ Ba(NO_3)_2(aq)\ +\ 2\ AgCl(s)$

Complete ionic:

$Ba^{2+}(aq)\ +\ 2\ Cl^-(aq)\ +\ 2\ Ag^+(aq)\ +\ 2\ NO_3^-(aq)\ \rightarrow\ Ba^{2+}(aq)\ +\ 2\ NO_3^-(aq)\ +\ 2\ AgCl\ (s)$

Net ionic:

$Ag^+(aq)\ +\ Cl^-(aq)\ \rightarrow\ AgCl(s)$

Example #4 (from Part B): $3\ NaOH(aq)\ +\ FeBr_3(aq)\ \rightarrow\ Fe(OH)_3(s)\ +\ 3\ NaBr(aq)$

Example #5: $H_2SO_4(aq)\ +\ 2\ NaOH(aq)\ \rightarrow\ Na_2SO_4(aq)\ +\ 2\ H_2O(l)$

Part D – Extra Practice

Write both the complete and net ionic equations for the balanced chemical reactions below.

a) $2\ KI(aq) + Pb(NO_3)_2(aq) \rightarrow 2\ KNO_3(aq) + PbI_2(s)$

b) $K_2SO_4(aq) + BaBr_2(aq) \rightarrow BaSO_4(s) + 2\ KBr(aq)$

c) $Na_2CO_3(aq) + Pb(NO_3)_2(aq) \rightarrow 2\ NaNO_3(aq) + PbCO_3(s)$

d) $Cr(NO_3)_3(aq) + K_3PO_4(aq) \rightarrow CrPO_4(s) + 3\ KNO_3(aq)$

e) $Li_2SO_4(aq) + SrCl_2(aq) \rightarrow 2\ LiCl(aq) + SrSO_4(aq)$

f) $2\ HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2\ H_2O(l)$

g) $Na_2CO_3(aq) + CoCl_2(aq) \rightarrow CoCO_3(s) + 2\ NaCl(aq)$

h) $Hg_2(NO_3)_2(aq) + 2\ KCl(aq) \rightarrow 2\ KNO_3(aq) + Hg_2Cl_2(s)$

i) $HCl(aq) + NH_4OH(aq) \rightarrow NH_4Cl(aq) + H_2O(l)$

j) $2\ HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2\ H_2O(l)$

k) $2\ NaC_2H_3O_2(aq) + Pb(NO_3)_2(aq) \rightarrow 2\ NaNO_3(aq) + Pb(C_2H_3O_2)_2(aq)$