Chemistry 20

Lesson 23 – Reactions in Solution

I. Writing Ionic Equations

Many chemical reactions can be represented by three different kinds of equations: non–ionic equations, total ionic equations and net ionic equations. For reactions in aqueous solution, the most correct are ionic equations since substances that are electrolytes undergo dissociation into ions. The ionic species in aqueous solution subsequently react as ions.

Non–ionic Equations

In non–ionic equations, the elements and compounds are written in their molecular or formula unit forms. For example:

$$2\ AgNO_{3\ (aq)} + BaCl_{2\ (aq)} \rightarrow 2\ AgCl_{\ (s)} + Ba(NO_3)_{2\ (aq)}$$

Total Ionic Equations

In total ionic equations, elements and compounds are written in the forms in which they are present in aqueous solution – electrolytes are written in ion form, and non–electrolytes, precipitates and gases are written in their molecular or formula unit forms.

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

Net Ionic Equations

In net ionic equations, only those molecules, formula units or ions that have changed in form (i.e., those that have participated in the chemical reaction) are included in the equation. Ions or molecules that do not change (**spectator species**) are omitted.

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

$2\ Ag^{+}_{\ (aq)} + 2\ Cl^{-}_{\ (aq)} \rightarrow 2\ AgCI_{\ (s)}$ which reduces to:

$$Ag^{+}_{\ (aq)} + Cl^{-}_{\ (aq)} \rightarrow AgCI_{\ (s)}$$

The following is a summary of rules to observe when writing net ionic equations,

1. Species that do not change or take part in the reaction are not shown.
2. The equations must be balanced, both in atoms and in electrical charge using the lowest whole number ratio.
3. Species are written in the form in which they actually exist.

II. Obtaining a Net Ionic Equation

The following examples illustrate the writing of net ionic equations.

Zinc reacts with hydrochloric acid.

$Zn_{(s)} + 2\ HCI_{(aq)} \rightarrow ZnCl_{2\ (aq)} + H_{2\ (g)}$ (non–ionic)

$Zn_{(s)} + 2\ H^{+}_{(aq)} + 2\ Cl^{-}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + 2\ Cl^{-}_{(aq)} + H_{2\ (g)}$ (total ionic)

$Zn_{(s)} + 2\ H^{+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2\ (g)}$ (net ionic)

Sodium reacts with water.

$2\ Na_{(s)} + 2\ HOH_{(l)} \rightarrow H_{2\ (g)} + 2\ NaOH_{(aq)}$ (non–ionic)

$2\ Na_{(s)} + 2\ HOH_{(l)} \rightarrow H_{2\ (g)} + 2\ Na^{+}_{(aq)} + 2\ OH^{-}_{(aq)}$ (total ionic)

$2\ Na_{(s)} + 2\ HOH_{(l)} \rightarrow H_{2\ (g)} + 2\ Na^{+}_{(aq)} + 2\ OH^{-}_{(aq)}$ (net ionic)

In some situations the net and total ionic equations are the same.

Zinc reacts with a copper (II) sulfate solution.

$Zn_{(s)} + CuSO_{4\ (aq)} \rightarrow ZnSO_{4\ (aq)} + Cu_{(s)}$ (non–ionic)

$Zn_{(s)} + Cu^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} + Cu_{(s)}$ (total ionic)

$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$ (net ionic)

Chlorine reacts with a potassium iodide solution.

$Cl_{2\ (g)} + 2\ K\ I_{(aq)} \rightarrow I_{2\ (s)} + 2\ KCl_{(aq)}$ (non–ionic)

$Cl_{2\ (g)} + 2\ K^{+}_{(aq)} + 2\ I^{-}_{(aq)} \rightarrow I_{2\ (s)} + 2\ K^{+}_{(aq)} + 2\ Cl^{-}_{(aq)}$ (total ionic)

$Cl_{2\ (g)} + 2\ I^{-}_{(aq)} \rightarrow I_{2\ (s)} + 2\ Cl^{-}_{(aq)}$ (net ionic)

A barium chloride solution reacts with a potassium sulfate solution.

$BaCl_{2\ (aq)} + K_2SO_{4\ (aq)} \rightarrow BaSO_{4\ (s)} + 2\ KCl_{(aq)}$ (non–ionic)

$Ba^{2+}_{(aq)} + 2\ Cl^{-}_{(aq)} + 2\ K^{+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4\ (s)} + 2\ K^{+}_{(aq)} + 2\ Cl^{-}_{(aq)}$ (total ionic)

$Ba^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4\ (s)}$ (net ionic)

III. Practice problems

Write the non–ionic, total ionic and net ionic equations for each of the following reactions. All reactions are in water solution.

1. Hydrochloric acid is added to a solution of barium hydroxide.

2. Magnesium metal is added to an aqueous solution of hydrogen bromide.

3. Calcium metal reacts with water.

4. Aqueous solutions of potassium sulfate and barium chloride are mixed.

5. An aqueous solution of washing soda, Na_2CO_3, is added to remove $Mg^{2+}_{(aq)}$ from water.

IV. Writing Net Ionic Equations – An Alternate Method

An alternate way of writing net ionic equations is to first list the reactant species that are present. In writing the species present, the following should be remembered.

1. Elements, molecular compounds, weak acids and ionic compounds of low solubility are written in non–ionic form.
2. Soluble ionic compounds and strong acids are written as separate aqueous ions.

From the list of reactant species present, those species that will react are selected and the net ionic equation is written. For example, consider the reaction between aqueous solutions of lead (II) nitrate and potassium iodide. The reactant species present are

$$Pb^{2+}_{(aq)} , NO_3^{-}{}_{(aq)} , K^{+}_{(aq)} , I^{-}_{(aq)}$$

From the cation–anion species present, it is evident that the only reaction possible is a precipitation reaction. The solubility table shows that the presence of $Pb^{2+}_{(aq)}$ and $I^{-}_{(aq)}$ ions in solution gives a compound of low solubility. Consequently, the equation for the reaction is written as

$$Pb^{2+}_{(aq)} + 2\ I^{-}_{(aq)} \rightarrow PbI_{2\ (s)}$$

which is the net ionic equation for the reaction.

V. Stoichiometry example

Example 1

A 600 mL sample of Calgary water contains dissolved calcium ions. When a solution containing sodium carbonate is added to the water sample a precipitate forms and is collected. If 3.74 mg of precipitate was collected, what was the concentration of the calcium ions in the sample?

The net ionic reaction is:

$$Ca^{2+}{}_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow CaCO_{3\,(s)}$$

$$c_{Ca^{2+}} = 3.74\times10^{-3}\,g(CaCO_3)\times\frac{1\,mol\,(CaCO_3)}{100.09\,g(CaCO_3)}\times\frac{1\,mol\,(Ca^{2+})}{1\,mol\,(CaCO_3)}\times\frac{1}{0.600\,L} = \mathbf{6.23\times10^{-5}\ mol/L}$$

VI. Assignment

For **all** of the following write the balanced non–ionic, total ionic, and net ionic equations.

1. Potassium metal reacts with water.
2. A lead (II) acetate solution reacts with a sodium sulfide solution to yield a precipitate.
3. Solutions of sodium sulfate and barium bromide are added together.
4. An aqueous solution of sodium carbonate is used to remove calcium ions from water.
5. A precipitate forms when potassium iodide is mixed with lead (II) nitrate.
6. A calcium nitrate solution is added to a sodium carbonate solution.
7. A precipitate forms when iron (III) nitrate reacts with sodium phosphate.

For questions 8 to 10 calculate the unknown.

8. A solution is known to contain $S^{2-}{}_{(aq)}$ ions. 100 mL of the sulfide solution was found to react with 58.0 mL of a 0.100 mol/L solution of lead (II) nitrate. What is the concentration of the sulfide ions in the solution?
9. Excess chlorine was bubbled through a 2.50 L solution in which the iodide ion concentration was 0.120 mol/L. What mass of iodide will be produced?
10. A water sample was found to contain a small amount of iron (III) ions. If 4.80 mL of 0.0200 mol/L sodium hydroxide was required to precipitate all the iron (III) ions from a 800 mL sample, what was the concentration of the iron (III) ions in the sample?