CHEMISTRY 101 **IONIC AND COVALENT COMPOUNDS**
NET IONIC EQUATIONS

INTRODUCTION

When a metal reacts with a nonmetallic element, the latter usually becomes negatively charged by the attraction of one or more electrons from the metal, which in turn becomes positively charged. In general each element tends to assume its most stable electron configuration, with its outer electron energy level saturated, as in the noble gases. Such electrically charged atoms are called ions. The bonding force between ions, due primarily to the attraction of unlike electrical charges, results in ionic bonding.

When two nonmetallic elements combine, however, they do so by mutual Sharing of pairs of electrons by both of the atomic nuclei. Such a bond is called a covalent bond. In many cases the electron pairs are not equally shared by the two nuclei, resulting in a polar covalent bond.

Most salts and many bases possess ionic bonding. The ions of such substances constitute the structural units in the crystalline solid,%, when melted by heat, the ionic lattice structure is broken down and the mixture of positive and negative ions is a strong electrical conductor. When ionic substances dissolve in water, the ions likewise become moving positive and negative particles, and the solutions are strong electrical conductors or electrolytes. Accordingly, in equations we write the formulas of such substances in solution or in the molten state as separate ions. Examples of strong electrolytes are NAOH, which supplies Na+ and OH- ions, and K_2SO_4, which supplies K+ and SO_4^2- ions. Although they can be regarded as strong electrolytes in that normally all the substance which is dissolved is present in the ionic form, salts and bases with low solubility will supply only a low concentration of ions to solution and therefore are weak conductors of electricity.

Substances with covalently bonded atoms generally have definite molecules as units in any state of matter. Many such substances, if soluble in water or other suitable solvent, give a mixture of electrically neutral molecules, and the solution is a nonconductor of electricity. Two examples of nonelectrolytes are sucrose, $C_{12}H_{22}O_{11}$, and acetone, CH_3COCH_3

Most acids and many bases possess polar covalent bonds. A few (strong) acids, such as HCl, when dissolves in the very polar solvent water, will result in almost complete ionization when the strong attraction of the positive proton of the acid for the negative pole of the water molecule and the attraction of the negative part of the acid for the positive pole of the water molecule results in the formation of hydrated protons, H^+ (aq), and hydrated anions of the acid, such as Cl (aq). These solutions are strong electrical conductors. When hydrogen chloride dissolves in a nonpolar solvent, such as one of the many organic liquids, the proton is not attracted away from the chloride ion by the solvent. Consequently no ions are present in these solutions, and they are not electrical conductors. Most (weak) acids and bases, although consisting of polar molecules, do not ionize to any appreciable extent, and dissolve largely through dipole-dipole interactions producing few ions. These solutions are only weak electrical conductors.

Ionic Equations

Equations for chemical reactions involving acids, bases, and salts are sometimes written with molecular formulas, as if actual molecules were the reacting species:

$$NaOH(aq) + HNO_3 (aq) \rightarrow NaNO_3(aq) + H_2O (l)$$

This is satisfactory when only the stoichiometric quantities of chemicals entering into a reaction are considered. As you now explore the experimental evidence for the reactions of such ionic substances, it will be more in accord with the facts for you to write equations using formulas for the principal ions or molecules as they exist in solutions.

Since data on electrical conductivity and on the chemical behavior of strong acids and bases and soluble salts indicate that the aqueous solutions of these substances consist of individual ions, equations for reactions between them will be written in terms of the principal substances (ions or molecules) actually present before and after a reaction. Thus the above equation is rewritten:

$$Na^+ + OH^- + H^+ + NO_3^- \rightarrow Na^+ + NO_3^- + HOH (l)$$

This equation is called a total ionic equation. Note that it is understood that the ions are hydrated in solution, even though the designation (aq) is omitted.

In this equation neither the Na^+ nor the NO_3^- has reacted (spectator ions). Both are present as separate ions before and after the reaction. They are therefore omitted, and the equation called a net ionic equation becomes:

$$OH^- + H^+ \rightarrow HOH (l)$$

Such an equation focuses attention on the essential changes: the disappearance of acid properties due to H^{+}' and of basic properties due to OH^-. The water formed is so slightly ionized that the concentration of these ions is too low to show their characteristic properties to any appreciable extent.

Ionic reactions occur whenever ions unite to form slightly ionized or insoluble substances. In the above neutralization of nitric acid by sodium hydroxide, the only factor causing the reaction is the great tendency of H^+ and OH- to unite to form the very slightly ionized water.

A typical example of ions uniting to form an insoluble substance (really a substance that is only slightly soluble) is expressed by the equation:

$$Ca^{2+} + CO_3{}^{2-} \rightarrow CaCO^3(s)$$

It. makes no difference whether the calcium ion is obtained from calcium chloride, calcium nitrate, or any other soluble calcium salt. Carbonate ion could obtained equally well from sodium carbonate, ammonium carbonate, or from any other soluble, highly ionized carbonate. The chemical change, as recorded by the net ionic equation, is the same in each case.

How to Write Net Ionic Equations

The techniques involved in writing net ionic equations involve learning a few simple rules of solubility and strength of electrolytes.

To write net ionic equations:

1. Write a balanced molecular equation for the reaction.
2. Write a total ionic equations for the reaction, using the solubility rules.
3. Write the net ionic equation by eliminating any common species (spectator ions) from both sides of the total ionic equation.

General Solubility Rules for Common Salts and Bases.

(1) all nitrates, and acetates are soluble ($AgC_2H_3O_2$ is only moderately soluble).

(2) all halides are soluble except silver, lead, and mercury.

(3) all sulfates are soluble except barium, lead, and (the slightly soluble) calcium & silver.

(4) all carbonates, phosphates, hydroxides, oxides, and sulfides are insoluble, EXCEPT potassium, sodium and ammonium salts. Moderately soluble exceptions are barium hydroxide, calcium hydroxide, and magnesium, calcium, strontium, and barium sulfides. Many hydrogen phosphate salts are soluble.

(5) all potassium, sodium, and ammonium salts are soluble.

(6) All silver salts are insoluble except nitrates, perchlorates, acetates, and sulfates.

Workshop on Net Ionic Equations Name:

Application of Principles

Using the solubility rules, write (a) the molecular equation, (b) the total ionic equation, and (c) the net ionic equation for following reactions (if any) taking place on mixing the following. Assume that enough water is present to dissolve all soluble substances. Indicate solids by (s) after the formula and soluble, weak electrolytes by (aq) after the formula.

1. Ammonium acetate and hydrochloric acid

   (a)

   (b)

   (C)

2. Cupric carbonate and nitric acid

   (a)

   (b)

   (C)

3. Magnesium hydroxide and hydrobromic acid

   (a)

   (b)

   (C)

4. Ammonium chloride and sodium nitrate

   (a)

   (b)

   (C)

5. Aluminum metal and dilute sulfuric acid

   (a)

   (b)

   (C)

6. Ammonium hydroxide and hydrogen sulfide

   (a)

   (b)

   (C)

7. Magnesium chloride and sodium carbonate
(a)

(b)

(C)

8. Nitric acid and magnesium acetate
(a)

(b)

(C)

9. Ammonium chloride and sodium hydroxide
(a)

(b)

(C)

10. Barium nitrite and calcium nitrate
(a)

(b)

(C)

11. Nitric acid and silver chloride
(a)

(b)

(C)

For each of the following, write only the <u>net</u> ionic equation.

12. Aluminum hydroxide and nitric acid

13. Ammonium hydroxide and sulfuric acid

14. Hydrogen sulfide and copper(II) acetate