NET IONIC EQUATIONS

Net ionic equations are useful in that they show only those chemical species directly participating in a chemical reaction. They are thus simpler than the overall equation, and help us to focus on the "heart" of the chemical change in a particular reaction. The keys to being able to write net ionic equations are the ability to recognize monatomic and polyatomic ions, the solubility rules, and the rules for electrolyte behavior. If you are weak in these areas, a review of these concepts would be helpful before attempting to write net ionic equations.

Electrolyte Behavior

When dissolved in water, many materials break apart into ions, or dissociate. This is because water is such a polar solvent. The main piece of experimental evidence for this is the fact that solutions of these materials will conduct an electrical current. Since an electrical current is a flow of charged particles, and water molecules are neutral, the only explanation for the flow of current is that the solute dissociated to produce the ions necessary for current flow. Any material whose aqueous solution will conduct an electrical current (i.e., which contains ions) is called an electrolyte.

Some materials are 100% dissociated into their ions in aqueous solution; these materials are termed strong electrolytes. Materials which dissociate only partially into their ions are termed weak electrolytes, and materials which do not dissociate at all are termed nonelectrolytes. (Note that the definitions of "strong" and "weak" electrolytes are not related to the overall concentration of the solution, but rather to the fraction of the material existing in dissociated form.) In general, weak electrolytes dissociate into ions only to a small extent (typically less than 5%).

When writing chemical equations for reactions occurring in aqueous solution, it is often useful to write them showing the actual species in solution (i.e., as ions or molecules, as appropriate), rather than using the full "molecular" formula for all reactants and products. How can one tell whether a given material is a strong, weak, or nonelectrolyte when dissolved in water? The following list summarizes the "rules" for electrolyte behavior.

1. All salts (ionic compounds) are strong electrolytes.

2. Most acids are weak electrolytes (i.e., they are "weak acids"). This generalization includes both inorganic and organic (i.e., carbon-containing, whose formulas usually contain C, H, and O) acids. The only common exceptions to this generalization are HCl, HBr, HI, HNO_3, $HClO_4$ and H_2SO_4. These six are strong acids.

3. Among bases, metal hydroxides are strong electrolytes (i.e., they are "strong bases"). Ammonia (NH_3) and organic bases (whose formulas usually contain C, H, and N, and which can be considered as being derived from ammonia by replacing one or more of the hydrogen atoms with carbon-containing groups) are weak.

4. All other materials are nonelectrolytes.

Appropriate application of the above rules is crucial for writing chemical equations in net ionic form. Examples of this process are given below.

Precipitation Reactions

Let's first start with a complete chemical equation and see how the net ionic equation is derived. For example, take the reaction of lead(II) nitrate with hydrochloric acid to form lead(II) chloride and nitric acid, shown below:

$$Pb(NO_3)_2\ (aq) + 2\ HCl\ (aq) \longrightarrow PbCl_2\ (s) + 2\ HNO_3\ (aq)$$

NET IONIC EQUATIONS

This complete equation may be rewritten in ionic form by using the solubility rules and rules for electrolyte behavior. All salts are strong electrolytes, therefore the lead(II) nitrate will be dissociated. Both hydrochloric acid and nitric acid are strong acids (they are on the list of exceptions) and will therefore be dissociated. The lead(II) chloride, however is insoluble—remember that all halides are soluble <u>except</u> the silver, lead, and mercury(I) halides. The above equation written in its dissociated ("ionic") form is:

$$Pb^{2+}(aq) + 2\ NO_3^{-}\ (aq) + 2\ H^{+}(aq) + 2\ Cl^{-}\ (aq) \longrightarrow PbCl_2\ (s) + 2\ H^{+}(aq) + 2\ NO_3^{-}\ (aq)$$

Notice that the stoichiometry of the balanced equation must be maintained. $Pb(NO_3)_2$ is dissociated into Pb^{2+} and **<u>2</u>** NO_3^{-}, and the 2 HCl molecules shown in the reactants are dissociated into **<u>2</u>** H^{+} and **<u>2</u>** Cl^{-}. At this point, one may cancel out those ions which have not participated in the reaction. Notice how the nitrate ions and hydrogen ions remain unchanged on both sides of the equation. Ions which don't change during the chemical reaction are called <u>spectator ions</u> and can be removed from the equation without destroying the equality (as long as they are removed in *exactly the same number* from both sides!).

$$Pb^{2+}(aq) + \cancel{2\ NO_3}^{-}\ (aq) + \cancel{2\ H}^{+}(aq) + 2\ Cl^{-}\ (aq) \longrightarrow PbCl_2\ (s) + \cancel{2\ H}^{+}(aq) + \cancel{2\ NO_3}^{-}\ (aq)$$

What remains is the <u>net</u> ionic equation, showing only those chemical species participating in the chemical process:

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

It is also possible to predict the net ionic equation given only the reactants. For example, suppose you had to determine the net ionic equation resulting from the mixing of solutions of barium bromide and sodium sulfate:

$$BaBr_2\ (aq)\ +\ Na_2SO_4\ (aq) \longrightarrow\ ?$$

One way to approach this problem is to determine what ions are in solution. Since both reactants are salts, they will be fully dissociated into Ba^{2+}, Br^{-}, Na^{+}, and SO_4^{2-}. We know that barium bromide is soluble, but will sodium ions or sulfate ions combine with barium ions to form an insoluble compound? Barium ions and sodium ions, both being positive in charge, will repel each other, so no compound is expected to form between them. Similarly, since bromide and sulfate ions both have a negative charge, we would expect no compound to form from this combination. On the other hand, sulfate ions and barium ions could easily form barium sulfate. Now it is just a matter of consulting the solubility rules to see if barium sulfate is soluble or insoluble. The solubility rule for sulfates is that they are all soluble except for those of strontium, barium, and lead; silver sulfate and calcium sulfate are partially soluble. As you can see from these rules, barium sulfate will be insoluble. The sodium ions must therefore combine with bromide ions to form sodium bromide. According to the solubility rules, sodium bromide should be soluble—all sodium salts are soluble, as are most halides. Now we can write a complete balanced equation:

$$BaBr_2\ (aq)\ +\ Na_2SO_4\ (aq) \longrightarrow BaSO_4\ (s)\ +\ 2\ NaBr\ (aq)$$

As before, the above equation can be rewritten in ionic form, showing the soluble species as ions in solution:

$$Ba^{2+}(aq) + 2\ Br^{-}\ (aq) + 2\ Na^{+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4\ (s) + 2\ Na^{+}(aq)\ +\ 2\ Br^{-}\ (aq)$$

Next, cross out the spectator ions:

$$Ba^{2+}(aq) + \cancel{2\ Br}^{-}\ (aq) + \cancel{2\ Na}^{+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4\ (s) + \cancel{2\ Na}^{+}(aq)\ +\ \cancel{2\ Br}^{-}\ (aq)$$

NET IONIC EQUATIONS

What remains is the balanced, net ionic equation:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4\,(s)$$

Next, let's consider the case when a solution of magnesium chloride is mixed with one of copper(II) nitrate:

$$MgCl_2\,(aq) + Cu(NO_3)_2\,(aq) \longrightarrow ?$$

As before, the most likely reaction is for the ions to "switch partners":

$$MgCl_2\,(aq) + Cu(NO_3)_2\,(aq) \longrightarrow Mg(NO_3)_2\,(?) + CuCl_2\,(?)$$

According to the solubility rules, both $Mg(NO_3)_2$ and $CuCl_2$ are soluble in water (all nitrates are soluble; most halides are soluble):

$$MgCl_2\,(aq) + Cu(NO_3)_2\,(aq) \longrightarrow Mg(NO_3)_2\,(aq) + CuCl_2\,(aq)$$

We now recognize that all four compounds are salts, and completely dissociated in water:

$$Mg^{2+}\,(aq) + 2\,Cl^{-}\,(aq) + Cu^{2+}\,(aq) + 2\,NO_3^{-}\,(aq) \longrightarrow$$

$$Mg^{2+}\,(aq) + 2\,NO_3^{-}\,(aq) + Cu^{2+}\,(aq) + 2\,Cl^{-}\,(aq)$$

When we go to remove spectator ions, we notice that all the ions on the reactant side are exactly the same as those on the product side. This would mean that we would cross out all the ions as spectator ions, leaving nothing behind! In fact, that's exactly what happens in this "reaction." Because no precipitate or weak electrolyte has formed, our (correct) prediction in this case would be "no reaction."

Acid-Base Reactions

Net ionic equations are often applied to acid-base reactions as well. The key to successfully writing the net ionic equation for acid-base reactions is to be able to distinguish between a strong and weak acid or base. The degree of dissociation for an acid is determined by the strength of the acid. For a strong acid, dissociation is complete (100%), thus we would show the acid dissociated into ions in the net ionic equation. In solution, weak acids dissociate to a small extent (usually < 5%). Since weak acids exist in solution predominantly as undissociated molecules, they are shown as neutral molecules (i.e., not shown to be dissociated into ions) in the net ionic equation. For the same reason, weak bases are also shown as undissociated in solution. Strong bases are shown dissociated into the metal cation and OH^{-}.

Organic acids usually contain C, H, and O in their molecular formula, and in particular, have a carboxylic acid functional group (COOH or CO_2H). All organic acids are weak. An example would be benzoic acid, $HC_6H_5CO_2$ or C_6H_5COOH. Ammonia (NH_3) and organic bases (usually derivatives of ammonia) are weak bases. For example methylamine, CH_3NH_2, is a weak base.

The products commonly produced by an acid-base reaction are a salt and water. To write the formula of the salt, it is helpful to know that the cation of the salt always comes from the base and the anion of the salt always comes from the acid. (One way to remember this is to keep the consonants together and the vowels together—**c**ation from the **b**ase and **a**nion from the **a**cid.)

NET IONIC EQUATIONS

Let's apply some of these rules to writing the net ionic equations for some acid-base reactions. What will be the complete, total ionic and net ionic equations for the reaction of an aqueous solution of nitric acid (HNO_3) with one of potassium hydroxide (KOH)?

$$HNO_3\ (aq)\ +\ KOH\ (aq) \longrightarrow\ ?$$

The cation of the base is potassium ion (K^+) and the anion of the acid is nitrate ion (NO_3^-), so the salt formed will be KNO_3 and the other product will be water. Since all nitrates are soluble, potassium nitrate will be aqueous, and the water will be in the liquid phase. The complete, balanced equation is:

$$HNO_3\ (aq)\ +\ KOH\ (aq) \longrightarrow KNO_3\ (aq)\ +\ H_2O\ (l)$$

According to the rules for electrolyte behavior given above, HNO_3, KOH, and KNO_3 are all strong electrolytes, and thus will dissociate completely. Water is in the "all other materials" category, so will be a nonelectrolyte and not shown dissociated. The total ionic equation will then be:

$$H^+\ (aq)\ +\ NO_3^-\ (aq)\ +\ K^+\ (aq)\ +\ OH^-\ (aq) \longrightarrow K^+\ (aq)\ +\ NO_3^-\ (aq)\ +\ H_2O\ (l)$$

We then eliminate the spectator ions (potassium and nitrate):

$$H^+\ (aq)\ +\ \cancel{NO_3}^-\ (aq)\ +\ \cancel{K}^+\ (aq)\ +\ OH^-\ (aq) \longrightarrow \cancel{K}^+\ (aq)\ +\ \cancel{NO_3}^-\ (aq)\ +\ H_2O\ (l)$$

What remains is the net ionic equation:

$$H^+\ (aq)\ +\ OH^-\ (aq) \longrightarrow H_2O\ (l)$$

This same net ionic equation is the result for the reaction of <u>any</u> strong acid with <u>any</u> strong base. (Try some other examples to convince yourself.)

Let's see what happens with the reaction of a weak acid with a strong base. What will be the complete, total ionic and net ionic equations for the reaction of benzoic acid ($HC_6H_5CO_2$) with sodium hydroxide?

The salt formed will be $NaC_6H_5CO_2$ (Na^+ from the base and $C_6H_5CO_2^-$ from the acid). This salt is soluble because all group IA compounds are soluble. Thus, the complete equation is:

$$HC_6H_5CO_2\ (aq)\ +\ NaOH\ (aq) \longrightarrow NaC_6H_5CO_2\ (aq)\ +\ H_2O\ (l)$$

Since benzoic acid is a weak acid (it contains carbon in its molecular formula, and is not one of the six common strong acids), it should <u>not</u> be shown dissociated in the total ionic or the net ionic equation. Both NaOH (metal hydroxide) and $NaC_6H_5CO_2$ (salt) are strong electrolytes. The total ionic equation therefore is:

$$HC_6H_5CO_2\ (aq)\ +\ Na^+\ (aq)\ +\ OH^-\ (aq) \longrightarrow Na^+\ (aq)\ +\ C_6H_5CO_2^-\ (aq)\ +\ H_2O\ (l)$$

In this case, sodium ions are the only spectator ions:

$$HC_6H_5CO_2\ (aq)\ +\ \cancel{Na}^+\ (aq)\ +\ OH^-\ (aq) \longrightarrow \cancel{Na}^+\ (aq)\ +\ C_6H_5CO_2^-\ (aq)\ +\ H_2O\ (l)$$

What remains is the net ionic equation:

$$HC_6H_5CO_2\ (aq)\ +\ OH^-\ (aq) \longrightarrow C_6H_5CO_2^-\ (aq)\ +\ H_2O\ (l)$$

NET IONIC EQUATIONS

Finally, let's look at an acid-base reaction involving a weak base and a strong acid. What will be the complete, total ionic and net ionic equations for the reaction of HCl with methylamine (CH_3NH_2), a weak organic base? Since hydroxide is not a part of the formula of the weak base, water will not be a product in this case. The complete reaction will be:

$$HCl\ (aq) + CH_3NH_2\ (aq) \longrightarrow CH_3NH_3Cl\ (aq)$$

HCl will be dissociated since it is a strong acid. Since methylamine is a weak base, it will be shown undissociated in the equation. Salts of organic bases are soluble (similar to NH_4^+) and will be dissociated. The total ionic equation that results is:

$$H^+\ (aq) + Cl^-\ (aq) + CH_3NH_2\ (aq) \longrightarrow CH_3NH_3^+\ (aq) + Cl^-\ (aq)$$

The only spectator ion present is chloride:

$$H^+\ (aq) + \cancel{Cl^-\ (aq)} + CH_3NH_2\ (aq) \longrightarrow CH_3NH_3^+\ (aq) + \cancel{Cl^-}\ (aq)$$

What remains is the net ionic equation:

$$H^+\ (aq) + CH_3NH_2\ (aq) \longrightarrow CH_3NH_3^+\ (aq)$$