Predicting the Products of Chemical Reactions and Writing Chemical Equations

Introduction

It is not always easy to predict the product of a chemical reaction. Often, a reaction must be carried out and the products analyzed in order to determine what formed. However, frequently a very good prediction can be made if one analyzes the type of reaction that occurs.

Concepts

- Chemical equations
- Reaction types

Background

Writing Chemical Equations

Chemical equations use symbols to show the reactants consumed in a reaction, and the products that form. There are several ways to write chemical equations. Each emphasizes different facets of the reaction. As an example, we will consider the mixing of two colorless solutions, sodium chloride and lead(II) nitrate, which forms a precipitate of lead(II) chloride and a solution of sodium nitrate. The balanced "molecular" equation for this reaction is:

$$2NaCl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$$

Here the symbol "aq" means aqueous solution and "s" means the formation of a solid precipitate. In reality, since a solution of sodium chloride consists of ions of sodium and chloride that are no longer attached to one another, writing the formula as NaCl is deceptive. The same is true for the lead(II) nitrate and sodium nitrate solutions. An equation that is more indicative of the substances actually present would be a "total ionic" equation:

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

The total ionic equation separates the positive and negative ions of substances whose ions are not attached to one another. In this equation you can see that the Na^+ ions and the NO_3^- ions do not really change. They appear as both reactants and products, and are called "spectator" ions. An equation that shows only the substances that actually change in the reaction is called a "net ionic equation." In a net ionic equation, spectator ions are not included:

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

Notice that in a balanced net ionic equation, both the atoms and the charges must balance.

When writing ionic equations, soluble ionic compounds that are strong electrolytes are written as separated ions. Since strong acids and bases ionize totally in water, their formulas are also written as separated ions. Weak acids and bases are written as molecular compounds. Even though a weak acid such as acetic acid does ionize slightly in water, the percent of ionization is very small. Only the formulas of the major species are shown in the net ionic equation. Solids, gases, and nonionic liquids, whether dissolved or not, are written as neutral, molecular formulas.

Types of Reactions

There are many different types of chemical reactions. If a chemist can identify the type of reaction that occurs, there is a good chance that the products of the reaction can be predicted. Some basic reaction types follow.

Combination Reaction. In this reaction, two or more simple substances combine to form a larger product. An example is heating sodium metal in chlorine gas to form solid sodium chloride:

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

Decomposition Reaction. In this reaction, a larger reactant breaks apart to form two or more simpler products. Often, heat, electricity, or a catalyst is needed to initiate the reaction. An example is the heating of solid sodium hydrogen carbonate to form sodium oxide, carbon dioxide, and water:

$$2NaHCO_3(s) \xrightarrow{\text{heat}} Na_2O(s) + 2CO_2(g) + H_2O(g)$$

Single Replacement. This is a reaction where an element reacts with a compound to form a new element and a different compound. An example is the reaction occurring when zinc metal is placed into a solution of copper(II) nitrate:

$$Zn(s) + Cu(NO_3)_2(aq) \rightarrow Cu(s) + Zn(NO_3)_2(aq)$$

The net ionic equation for this reaction would be:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

Since nitrate ions are spectator ions, they are not included.

Double Replacement (also called **Exchange** or **Metathesis Reactions**). In this type of reaction, two water-soluble ionic compounds form a new substance, composed of atoms from both of the initial compounds. The new substance does not dissociate in solution, but may be a solid that precipitates out of solution, a gaseous substance that bubbles out, or a new molecular compound that is soluble. An example is the reaction of aqueous solutions of strontium nitrate and potassium hydroxide to form a precipitate of strontium hydroxide and a solution of potassium nitrate:

$$Sr(NO_3)_2(aq) + 2KOH(aq) \rightarrow Sr(OH)_2(s) + 2KNO_3(aq)$$

The net ionic equation is:

$$Sr^{2+}(aq) + 2OH^-(aq) \rightarrow Sr(OH)_2(s)$$

Oxidation–Reduction (Redox Reactions). In this type of reaction, electrons are transferred from one substance to another. There must be both a substance oxidized and a substance reduced. An example is the reaction between liquid bromine and a solution of sodium iodide forming iodine and aqueous sodium bromide:

$$Br_2(l) + 2\,NaI(aq) \rightarrow I_2(aq) + 2NaBr(aq)$$

In this reaction, bromine is reduced and iodide ion is oxidized. This is also a single replacement reaction. The net ionic equation is:

$$Br_2(l) + 2I^-(aq) \rightarrow I_2(aq) + 2Br^-(aq)$$

Acid–Base. In acid–base reactions, hydrogen ions (protons) are exchanged between two substances. One substance must gain hydrogen ions, and another must lose them. An example is the addition of concentrated hydrochloric acid to a solution of sodium sulfide:

$$2HCl(aq) + Na_2S(aq) \rightarrow H_2S(g) + 2NaCl(aq)$$

This reaction is also a double replacement reaction. Notice that the hydrogen ion is transferred to the sulfide ion. Since hydrochloric acid is a strong acid and hydrosulfuric acid is a weak acid, the net ionic equation is:

$$2H^+(aq) + S^{2-}(aq) \rightarrow H_2S(g)$$

Complex Formation. Complex formation is a special type of combination reaction where a substance known as a coordination compound or complex ion forms. An example is the reaction that occurs when concentrated ammonia is added to a solution of copper(II) nitrate, forming an aqueous solution of tetraamminecopper(II) nitrate:

$$4NH_3(aq) + Cu(NO_3)_2(aq) \rightarrow [Cu(NH_3)_4](NO_3)_2(aq)$$

The net ionic equation is:

$$4NH_3(aq) + Cu^{2+}(aq) \rightarrow [Cu(NH_3)_4]^{2+}(aq)$$

Combustion. Combustion, or burning, always requires the addition of oxygen. In a combustion reaction, all of the elements in the compound combine with oxygen. Carbon dioxide and water are common products of the combustion of a compound containing carbon and hydrogen. An example is the burning of ethanol in air:

$$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

It is important to note that many reactions fit into more than one category. For example, many single replacement reactions are also oxidation–reduction reactions. Acid–base reactions may also be double replacement reactions.

In order to predict the products of a reaction, it is very helpful to first identify the type of reaction that is occurring.

The Advanced Placement Exam

The Advanced Placement Chemistry Exam recognizes that two of the necessary skills for a chemist are the ability to predict the products of a chemical reaction, and the ability to write a balanced chemical equation that describes that reaction. In a required, free-response question, students are asked to predict products of a chemical reaction and to write a net ionic equation for the process. It is not necessary to balance the equation because that skill is tested in other questions.

Helpful Hints for Writing Equations and Predicting Products

1. Learn the formulas of common ions and their charges. Formulas and charges must be written correctly.
2. Write formulas of separated ions with their proper charges. If an ion is written without a charge, it is incorrect. Solids, gases, and molecular compounds are written as neutral substances.
3. Know the solubility rules for ionic compounds. These are necessary to predict the formation of a precipitate.
4. Learn which acids and bases are strong and which are weak. Strong acids and bases are strong electrolytes, and are written as separated ions. Weak acids and bases are written as neutral molecules.
5. Try to classify the type of reaction: combination, single replacement, acid–base, precipitate formation, etc.
6. When a free element is a reactant or product, the reaction must be an oxidation–reduction reaction. There must be both a substance oxidized and a substance reduced.
7. Learn some common oxidizing agents and their products:

   $$MnO_4^- \rightarrow Mn^{2+} \qquad CrO_4^{2-} \rightarrow Cr^{3+} \qquad Cr_2O_7^{2-} \rightarrow Cr^{3+} \qquad NO_3^- \rightarrow NO$$

   In order to complete these half-reactions, you must add H^+ and H_2O. Suspect this type of reaction when "acidified" conditions are specified.
8. Nonmetal oxides are "acid formers" and metal oxides are "base formers." Nonmetal oxides plus water form acids, metal oxides plus water form bases. The nonmetal or metal keeps the same oxidation number. Nonmetal oxides plus bases form a salt and water. Metal oxides plus acids form a salt and water. Nonmetal oxides plus metal oxides form salts.
9. The decomposition of a carbonate forms an oxide and carbon dioxide gas.

   $$BaCO_3 \rightarrow BaO + CO_2$$

   The decomposition of a hydroxide forms an oxide and water.

   $$Ca(OH)_2 \rightarrow CaO + H_2O$$

10. Amphoteric ions such as Zn^{2+} and Al^{3+} form hydroxides that dissolve in both acid and base. In basic solution you may find $Zn(OH)_4^{2-}$ and $Al(OH)_4^-$ or $Al(OH)_6^{3-}$. Hydrogen ions from acid may neutralize some of the hydroxides, precipitating the neutral hydroxide compound, or all of the hydroxide, leaving the metal ion in solution.
11. Ammonia forms complex ions such as $Ag(NH_3)_2^+$, $Cu(NH_3)_4^{2+}$, $Cd(NH_3)_4^{2+}$, $Zn(NH_3)_4^{2+}$. These are broken up by addition of acid. The hydrogen ion from the acid changes the NH_3 to NH_4^+ which can no longer bond to the metal ion.
12. For partial credit, write down formulas for reactants even if you don't know products. If you know some of the products, write their formulas.
13. Even though the equations need not be balanced, every atom included in the reaction must appear in some form on both sides of the equation.
14. It is not necessary to include symbols for the state (solid, liquid, gas, aqueous). Do not include the spectator ions. Do not take the time to balance the equations. Do not answer more than asked for.

Examples of Writing Equations and Predicting Products

Write net ionic equations showing the reactants and products for the following reactions. Assume reactions occur in aqueous solution unless otherwise indicated. Do not include formulas for ionic or molecular species that do not take part in the reaction. It is not necessary to balance the equations or to include symbols for the state. A reaction occurs in all cases.

1. **Calcium metal is burned in air.**

   The reactive element in air is oxygen. Since this reaction begins with two simple substances, it must be a combination reaction. It is also a redox equation; calcium is oxidized and oxygen is reduced. The unbalanced equation is:

$$Ca + O_2 \rightarrow CaO$$

2. **Chlorine gas is bubbled into a solution of potassium iodide.**

   Here the reactants are an element, chlorine, and a compound, potassium iodide. We can predict a single replacement reaction. Since chlorine is a nonmetal, it will replace the negative ion. The unbalanced ionic equation is:

$$Cl_2 + K^+ + I^- \rightarrow I_2 + K^+ + Cl^-$$

   Since a net ionic equation is wanted, the potassium ions must be dropped:

$$Cl_2 + I^- \rightarrow I_2 + Cl^-$$

   This is also an oxidation–reduction equation which can be predicted because a neutral element reacts. The chlorine is reduced and the iodide ion is oxidized.

3. **Solid calcium carbonate is heated strongly.**

   This must be a decomposition reaction. Carbonates decompose to release carbon dioxide gas.

$$CaCO_3 \rightarrow CaO + CO_2$$

4. **Equal moles of hydrochloric acid and sodium phosphate are mixed in aqueous solution.**

   This is an acid–base reaction, a hydrogen ion will be transferred. Hydrochloric acid is a strong acid, so it is written as separate ions. The salt sodium phosphate is also written as separate ions. The unbalanced total ionic equation is:

$$H^+ + Cl^- + Na^+ + PO_4^{3-} \rightarrow HPO_4^{2-} + Na^+ + Cl^-$$

The ion HPO_4^{2-} forms because equal moles of acid and phosphate were mixed. To obtain the net ionic equation, the spectator ions Na^+ and Cl^- must be removed:

$$H^+ + PO_4^{3-} \rightarrow HPO_4^{2-}$$

5. **Potassium permanganate solution is added to an acidified solution of iron(II) nitrate.**

This is an example of a redox reaction. Permanganate ion is an oxidizing agent. The manganese is reduced to Mn^{2+}. Iron(II) will be oxidized to iron(III). Hydrogen ions and water must be included to account for all the atoms. The unbalanced total ionic equation is:

$$K^+ + MnO_4^- + Fe^{2+} + NO_3^- + H^+ \rightarrow Mn^{2+} + Fe^{3+} + K^+ + NO_3^- + H_2O$$

Eliminating the spectator ions leaves:

$$MnO_4^- + Fe^{2+} + H^+ \rightarrow Mn^{2+} + Fe^{3+} + H_2O$$

6. **Solid potassium oxide is added to water.**

Potassium oxide, a metal oxide, is a base former. Since it is a solid, its formula is written as a neutral compound and not separated into ions:

$$K_2O + H_2O \rightarrow K^+ + OH^-$$

7. **Solutions of sodium chromate and silver nitrate are mixed.**

The initial solutions are both soluble, ionic substances. The solubility rules predict that silver chromate is insoluble, and sodium nitrate is soluble. This is a double replacement (metathesis) reaction. The unbalanced total ionic equation is:

$$Na^+ + CrO_4^{2-} + Ag^+ + NO_3^- \rightarrow Ag_2CrO_4 + Na^+ + NO_3^-$$

Eliminating the spectators leaves:

$$CrO_4^{2-} + Ag^+ \rightarrow Ag_2CrO_4$$

8. **A solution of copper(II) sulfate is electrolyzed between graphite electrodes.**

This must be an oxidation–reduction equation. Graphite is inert, so will not take part. Copper(II) will be reduced to copper metal. An oxidation must also occur. The sulfate ion contains sulfur in the +6 oxidation state, so it cannot be oxidized any higher. It must be the water that is oxidized. Oxygen in water has a –2 oxidation state; it can be oxidized to neutral oxygen gas. The unbalanced equation is:

$$Cu^{2+} + H_2O \rightarrow Cu + O_2 + H^+$$

Experiment Overview

In this experiment you will carry out a number of chemical reactions, describe the reactants, predict the products, and describe what actually occurs. Then you will write a net ionic equation for the reaction. Your teacher will indicate which reactions you should actually carry out, and those for which you should only predict the products and write the equations.

Pre-Lab Questions

1. Use solubility rules to predict whether the following substances are soluble or insoluble in water:

   K_2CrO_4 ______________ $Cu(OH)_2$ ______________ Ag_2CO_3 ______________

   $NiSO_4$ ______________ $Pb(C_2H_3O_2)_2$ ______________ Al_2S_3 ______________

2. In the following equation identify the substances oxidized and reduced:

   $$K_2Cr_2O_7(aq) + 3SnCl_2(aq) + 14HCl(aq) \rightarrow 2CrCl_3(aq) + 3SnCl_4(aq) + 7H_2O(l) + 2KCl(aq)$$

   Oxidized: ______________ Reduced: ______________

   Write the above equation as a total ionic equation:

   Write the above equation as a net ionic equation:

3. Consult an activity series or a table of reduction potentials to predict if the following single replacement reactions will occur in aqueous solution:

   $Al(s) + 3NaCl(aq) \rightarrow 3Na(s) + AlCl_3(aq)$ ______________

   $Cl_2(g) + 2NaBr(aq) \rightarrow Br_2(l) + 2NaCl(aq)$ ______________

   $Co(s) + NiCl_2(aq) \rightarrow Ni(s) + CoCl_2(aq)$ ______________

4. Write formulas for the following substances as they would appear in a total ionic equation:

   aqueous calcium chloride ______________

   aqueous sulfuric acid ______________

   gaseous hydrogen sulfide ______________

   aqueous acetic acid ______________

   solid lithium carbonate ______________

   aqueous ammonia ______________

   aqueous sodium dihydrogen phosphate ______________

Materials

Chemicals

Refer to experimental directions for chemicals needed in each reaction.

Equipment

Alligator clips and wires
Battery, 9-volt
Beaker, 50-mL
Beral-type pipets
Bunsen burner
Centrifuge (optional)
Crucible
Crucible triangle
Electrodes, graphite (pencil lead)
Graduated cylinder, 10-mL
Litmus paper
Ring stand and ring
Spatula
Stirring rods
Test tubes, small
Tongs

Safety Precautions

Concentrated ammonia solution, concentrated hydrochloric acid solution, and sodium hydroxide solution are toxic by inhalation, ingestion, and is corrosive to all body tissues. Barium hydroxide solution is toxic by ingestion. Calcium oxide is a corrosive material and a severe body tissue irritant. Avoid all body tissue contact. Reaction of calcium oxide and water will produce large amounts of heat and skin burns are possible. A lump of calcium oxide may disintegrate violently and splatter when water is added. Copper(II) sulfate solution is slightly toxic by ingestion. Ethyl alcohol solution is flammable and a fire risk. Keep away from open flames and other sources of ignition. The addition of denaturants makes the 95% ethyl alcohol solution poisonous; it cannot be made non-poisonous. Iron(III) nitrate may be a skin and body tissue irritant. Hydrogen peroxide solution is an oxidizer and skin and eye irritant. Magnesium is a flammable solid; it burns with an intense flame. Manganese dioxide is a body tissue irritant and a strong oxidant; avoid contact with organic substances. Oxalic acid solution is slightly toxic by ingestion and is a skin and eye irritant. Potassium thiocyanate solution is slightly toxic by ingestion and emits fumes of cyanide in contact with concentrated acids. Silver nitrate solution is moderately toxic by ingestion. Solution will stain skin and clothing. Sodium phosphate solution is a body tissue irritant. Sulfuric acid solution is severely irritating to eyes and skin and slightly toxic by ingestion and inhalation. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory.

Procedure

Carry out the reactions as described. Create a Data Table in which you describe the reactants, products, and any indication that a reaction has occurred. Identify the type of reaction. Write a net ionic equation for the reaction. It is not necessary to balance the equation.

1. Solutions of sodium chloride and silver nitrate are combined.

   Mix 1 mL each of 0.1 M sodium chloride solution and 0.1 M silver nitrate solution in a small test tube. Save the products for step 2.

2. Excess ammonia is added to solid silver chloride.

   Allow the precipitate from step 1 to settle (or centrifuge). Use a Beral-type pipet to remove the supernatant liquid and discard it. Add 1 mL of 6 M ammonia to the precipitate and stir. Save the solution for Part 3.

3. Excess hydrochloric acid is added to a solution of diamminesilver(I) chloride.

   Add 2 mL 6 M hydrochloric acid to the solution from step 3. Test with litmus to be sure the solution is acidic. If the solution is not acidic, add 1 more mL of 6 M hydrochloric acid.

4. Ethyl alcohol is burned.

   Place 0.5 mL ethanol in a crucible and ignite.

5. Calcium oxide powder is added to distilled water.

   Add the tip of a spatula of solid calcium oxide to distilled water. Test the resulting solution with litmus paper.

6. A solution of potassium iodide is electrolyzed using inert electrodes.

   Put 10 mL of 0.1 M potassium iodide solution in a small beaker. Using a 9-volt battery as a power source and graphite electrodes, allow an electric current to pass through the solution for two minutes. Test the solution near the electrodes with litmus paper.

7. Magnesium metal is burned in air.

   Hold a 1.5 cm-long strip of magnesium with a tongs, and ignite it in the flame of a Bunsen burner. *Do not look directly at the burning metal.*

8. Acetic acid solution is added to solid sodium hydrogen carbonate.

   Place a small spatula of sodium hydrogen carbonate in a test tube. Add 1 mL of 0.1 M acetic acid.

9. Potassium permanganate solution is added to acidified oxalic acid solution.

   Place 2 mL of oxalic acid solution in a test tube, add three drops of 0.1 M sulfuric acid, stir, and then add five drops of 0.1 M potassium permanganate solution.

10. A piece of metallic aluminum is dropped into a solution of copper(II) chloride.

    Place 2 mL of 0.1 M copper(II) chloride in a test tube. Make a loose ball of a small piece of aluminum foil (2-cm square) and drop it into the solution. Use a stirring rod to push it under the solution. Observe after five minutes.

11. Excess hydrochloric acid is added to a solution of sodium phosphate.

    Place 1 mL of 0.1 M sodium phosphate in a test tube, and add 5 mL of 0.1 M hydrochloric acid. Feel the outside of the test tube to see if there is evidence of a reaction taking place.

12. Magnesium metal is placed in a solution of hydrochloric acid.

    Place 2 mL of 1.0 M hydrochloric acid in a test tube. Add a 1 cm long strip of magnesium metal.

13. Solid sodium hydrogen carbonate is added to sulfuric acid solution. Place 2 mL of 0.1 M sulfuric acid in a test tube. Add a small spatula of solid sodium hydrogen carbonate.

14. Potassium thiocyanate solution is added to a solution of iron(III) nitrate.

    Place 2 mL of 0.1 M iron(III) nitrate in a test tube. Add several drops of 0.1 M potassium thiocyanate solution.

15. Excess sodium hydroxide solution is added to a solution of aluminum nitrate.

    Place 2.0 mL of 0.1 M aluminum nitrate in a test tube. Dropwise, with stirring, add 2 mL of 1.0 M sodium hydroxide.

16. Solutions of sodium hydroxide and copper(II) sulfate are combined.

    Place 1 mL of 0.1 M copper(II) sulfate in a test tube. Add 3 mL of 0.1 M sodium hydroxide solution.

17. A solution of ammonia is added to a solution of iron(III) nitrate.

    Place 2.0 mL of 0.1 M iron(III) nitrate in a test tube. Add 2.0 mL of 0.1 M ammonia solution.

18. Excess sulfuric acid is added to a solution of barium hydroxide.

    Place 1.0 mL of 0.1 M barium hydroxide in a test tube. Add 5 mL of 0.1 M sulfuric acid.

19. Copper wire is dropped into a solution of silver nitrate.

    Place 5 mL of 0.1 M silver nitrate in a test tube. Add a small coil of copper wire. Observe after 5 minutes.

20. Solid manganese(IV) oxide is dropped into a solution of hydrogen peroxide.

    Place 5 mL of 3% hydrogen peroxide in a test tube. Add a small amount (the tip of a spatula) of solid manganese(IV) oxide.

Disposal and Cleanup

Your teacher will provide disposal and cleanup instructions.

Data and Analysis

For each of the reactions, give formulas and descriptions of reactants and products, evidence of a chemical reaction, identify the reaction type, and write a net ionic equation. It is not necessary to balance the equation. Some reactions can be categorized as more than one type of reaction.

You may find it helpful to prepare a chart for each reaction such as the one shown below. Also, you may find it easier to write a total ionic equation first, and then cancel out spectator ions.

1. Solutions of sodium chloride and silver nitrate are combined.

Formulas and Description of Reactants	Formulas and Description of Products	Evidence of a Chemical Reaction

Type of Chemical Reaction	Unbalanced Net Ionic Equation