

EXPERIMENT 5 – Double Replacement Reactions

PURPOSE

- To identify the ions present in various aqueous solutions.
- To systematically combine solutions and identify the reactions that form precipitates and gases.
- To become familiar with writing equations for reactions, including net ionic equations.

EQUIPMENT & CHEMICALS

Spot plates (or small test tubes)

Stirring rod or toothpicks

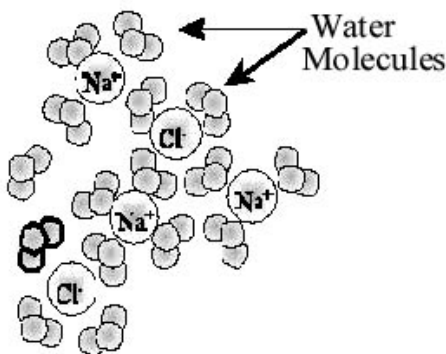
0.1 M solutions of the following:

| | | | | | | |
|-----------------------------------|-----------------------------------|--|-----------------------------------|---|-------------------|---------------------------------|
| AgNO ₃ | Ca(NO ₃) ₂ | CuSO ₄ | Fe(NO ₃) ₃ | HC ₂ H ₃ O ₂ | HCl | K ₂ CrO ₄ |
| KI | NaCl | NaC ₂ H ₃ O ₂ | Na ₂ CO ₃ | NaOH | Na ₂ S | NH ₄ Cl |
| Pb(NO ₃) ₂ | | | | | | |

INTRODUCTION

When one substance dissolves in another substance, a **solution** is formed. A **solution** is a homogeneous mixture in which the components are uniformly mixed. A solution consists of **solute** (the species that is dissolved) and **solvent** (the medium in which the solute has dissolved). The solvent is usually present in larger amount than the solute. When water is the solvent, the solution is called an **aqueous solution**.

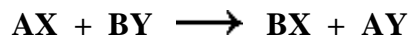
When an ionic compound dissolves in water, it dissociates into its constituent ions. Such a compound is a strong electrolyte and conducts electricity well in dilute aqueous solutions. For example, when NaCl dissolves in water, it dissociates into separate Na⁺ and Cl⁻ ions.



This process occurs as polar water molecules orient themselves around the sodium and chloride ions and pull them free from the solid crystal. Once removed from the solid crystal, the ions remain separated and surrounded by water molecules. There are no solid NaCl particles present, although in the solution, occasional contact between the Na⁺ and Cl⁻ ions, called **ion pairing**, does occur.

The ability of a substance to dissolve is called its **solubility**. The solubility of a substance in water is typically defined as the mass of the substance that will dissolve in 100 mL of water. Solubility is temperature-dependent and usually increases for solid solutes dissolving in a liquid solvent.

Double replacement reactions (also called “double displacement” or “exchange” or “metathesis” reactions) have the general form



Double replacement reactions typically form a product that is either molecular or ionic. Molecular products such as H₂O remain in solution and may not appear visually, but gaseous molecular substances such as CO₂ are usually identified easily by the appearance of bubbles and/or a new odor.

A chemical reaction in which an insoluble product (or precipitate) forms is called a **precipitation reaction**. The reactants are usually soluble, but the product formed is insoluble and separates out as a solid.

Reactions in aqueous solution can be written three ways:

- 1) as the **formula** or “**molecular**” equation
- 2) as the **complete ionic** equation
- 3) as the **net ionic** equation

A **formula equation** uses the normal formulas of the reactants and products. The **complete ionic equation** represents the formulas of soluble ionic compounds and strong electrolytes more realistically as their separated positive and negative ions in solution. The **net ionic equation** is a simplified ionic equation in which the same ions appearing on both sides are omitted. Such ions are called **spectator ions** because they undergo no chemical change themselves; they are only acting as “spectators” to the reaction. The net ionic equation contains the ions that do not “cancel” as spectator ions and products which are molecular substances (nonelectrolytes such as water or a gas or a weak electrolyte such as acetic acid, HC₂H₃O₂), or as insoluble solids (such as PbI₂). If no substances or ions remain in the net ionic equation (all of the ions are spectators and cancel!) then there is no net reaction.

Guidelines for Writing Formula Equations

Step 1. Write the reactants on the left side of your equation, taking care that their **formulas** are correct.

Step 2. Following the pattern $\text{AX} + \text{BY} \rightarrow \text{BX} + \text{AY}$, complete the reaction by writing the formulas of the products on the right side of your equation. Be sure all of the formulas are correct at this stage!

Step 3. Next, **balance** your equation using **coefficients** in front of the formulas. The coefficients in a properly balanced equation are the lowest possible whole numbers.

Step 4. Now, indicate the **physical state** of each substance in your equation, (s), (l), (g), or (aq) for solid, liquid, gas, or aqueous solution (dissolved). In this lab, each reactant substance is already in aqueous solution, so you can automatically indicate this using the “(aq)” label.

How can you determine the physical states of the products? Your observations of the reactions will provide valuable clues. If the mixture becomes cloudy due to the formation of tiny insoluble solid particles which eventually settle or “precipitate” to the bottom of the reaction vessel, you know that a solid product formed (s).

Which substance in your equation specifically is the solid? We refer to a table of **solubility rules** to determine this (see below).

If you observe the formation of bubbles, or notice a new odor, you know that one of the products is a gaseous substance (g). Some gaseous products are CO₂, NH₃, and H₂S.

SOLUBILITY RULES

One of the factors driving a double replacement reaction is the formation of a solid precipitate. A precipitate is an insoluble solid compound formed during a chemical reaction in solution. To predict whether a precipitate will form when you mix together two ionic reactants, you need to know whether any of the possible products are insoluble. Considering the number of ionic compounds, it would be very difficult to memorize the solubilities of so many compounds. Fortunately we can group compounds into solubility categories. This is done with a set of rules called **solubility rules**.

Solubility Rules

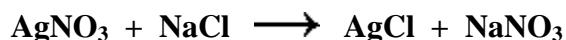
| | Rule Statement | Exceptions |
|----|--|---|
| 1 | All Group 1A and ammonium (NH ₄ ⁺) compounds are soluble. | -- |
| 2 | All nitrates (NO ₃ ⁻) are soluble | -- |
| 3 | Most acetates are soluble. | AgC ₂ H ₃ O ₂ * |
| 4 | Most chlorides, bromides, and iodides are soluble. | AgCl, Hg ₂ Cl ₂ , PbCl ₂ ,* AgBr, HgBr ₂ , Hg ₂ Br ₂ , PbBr ₂ ,* AgI, HgI ₂ , Hg ₂ I ₂ , PbI ₂ |
| 5 | Most sulfates are soluble. | CaSO ₄ , SrSO ₄ , BaSO ₄ , Ag ₂ SO ₄ , Hg ₂ SO ₄ , PbSO ₄ |
| 6 | Most carbonates are <u>insoluble</u> . | Group 1A carbonates, (NH ₄) ₂ CO ₃ |
| 7 | Most phosphates are <u>insoluble</u> . | Group 1A phosphates, (NH ₄) ₃ PO ₄ |
| 8 | Most sulfides are <u>insoluble</u> . | Group 1A sulfides, (NH ₄) ₂ S |
| 9 | Most hydroxides are <u>insoluble</u> . | Group 1A hydroxides, Sr(OH) ₂ ,* Ba(OH) ₂ * |
| 10 | Most chromates are <u>insoluble</u> . | Group 1A chromates, (NH ₄) ₂ CrO ₄ |

* Moderately soluble

Example 1 – Formation of a Solid Product

Let us consider the reaction of an aqueous solution of NaCl with an aqueous solution of AgNO₃. When we place a few drops of the NaCl solution in the reaction container followed by a few drops of AgNO₃ solution, we observe an immediate cloudiness (white precipitate) which indicates that a solid product has formed. A precipitation chemical reaction has occurred.

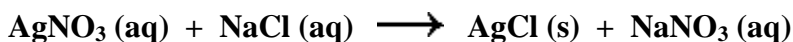
Following Steps 1-3 given above, the balanced equation for the reaction is,



Now we are ready for Step 4, the identification of the physical states of each substance. Referring to the solubility rules, we determine that AgNO₃ is soluble (aq) because all nitrates are soluble (Rule 2). NaCl is also soluble because all Group 1A compounds are soluble (Rule 1).

Looking at the products, we see from the solubility rules that AgCl is insoluble (Rule 4 exception) and NaNO₃ is soluble (Rules 1 and 2). Therefore, the white precipitate that was observed is identified as silver chloride, AgCl.

Adding the labels for the physical states, our reaction is now written,



Guidelines for Writing Complete Ionic Equations

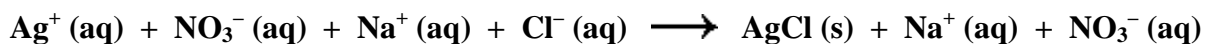
The **complete ionic equation** shows which reactants and products exist as separate positive and negative ions in solution and which do not. The steps for writing the complete ionic form of a reaction are as follows:

Step 1. First write the balanced formula equation following Steps 1-4 above. In the example above, this is the reaction,



Step 2. All soluble ionic compounds are separated into their positive and negative ions. These are the ionic compounds designated “(aq)” in the reaction. Molecular substances that are strong electrolytes (in particular, strong acids such as HCl and HNO₃) are also separated into their positive and negative ions in aqueous solution.

After Step 2, the above reaction becomes,



As you can see, this is a much lengthier, but more correct, way of writing the reaction.

Note that the solid product, AgCl, was not separated into its positive and negative ions. In the solid state these ions are not separated from each other, so we do not separate them in the equation either. This is always the case for solid reactants or products.

The formulas of molecular substances such as H₂O and CO₂ are also not separated into positive and negative ions, for the same reason as with solids: these substances do not exist as separated positive and negative ions. There are some exceptions as noted above: strong acids such as HCl and HNO₃.

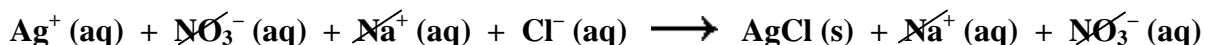
A final category of substance whose formulas are not separated into positive and negative ions are **weak electrolytes**. In aqueous solution, weak electrolytes separate into positive and negative ions only to a very small degree, and as such are essentially molecular in nature. An example is the weak acid acetic acid, HC₂H₃O₂, which is the acid in vinegar.

Guidelines for Writing the Net Ionic Equation

This is a simplified equation for the reaction that omits ions that are “spectators.” Spectator ions simply remain in solution before and after the reaction. As such, they do not combine with another ion to form an insoluble solid, molecular, or weak electrolyte product.

Step 1. Write the complete ionic equation as outlined above.

Step 2. “Cancel” the ions that appear on both sides of the equation. In the above example, these “spectator” ions are crossed out:



This leaves the net ionic equation,

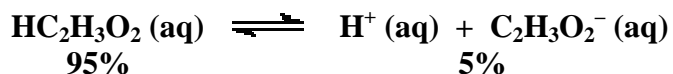


The spectator ions in this reaction are sodium ion, Na^+ , and nitrate ion, NO_3^- . They remained in solution without forming any new product.

The ions that do participate in the formation of the new product are called participating ions. In the above reaction, Ag^+ and Cl^- are the participating ions.

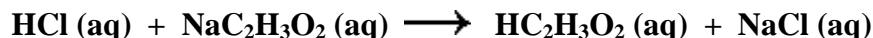
Example 2 – Formation of a Weak Electrolyte

Weak electrolytes dissociate only to a small degree. Typically, an aqueous solution of a weak electrolyte such as acetic acid contains mostly undissociated molecules with only 5% or less in dissociated form:

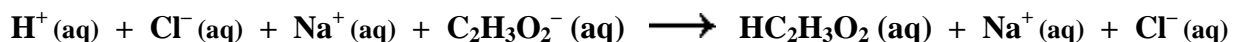


Following is the reaction of hydrochloric acid (a strong acid) with sodium acetate:

Formula Equation:

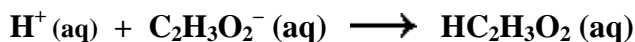


Complete Ionic Equation:



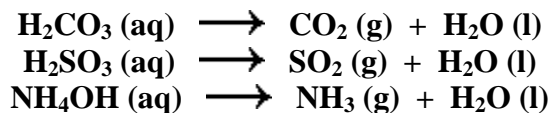
Hydrochloric acid, $\text{HCl} (\text{aq})$, being a strong acid, is written in its dissociated form, $\text{H}^+ (\text{aq}) + \text{Cl}^- (\text{aq})$. However, acetic acid, being a weak acid, is written in its undissociated, molecular form, $\text{HC}_2\text{H}_3\text{O}_2 (\text{aq})$.

Net Ionic Equation:

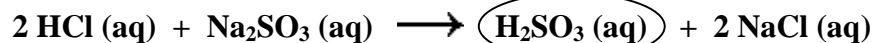


Products that Decompose Upon Formation

Certain products such as H_2CO_3 (carbonic acid), H_2SO_3 (sulfurous acid), and NH_4OH (ammonium hydroxide) are unstable in aqueous solution and immediately decompose according to the following reactions:



When these products form, you should write the decomposition products in your reactions. For example, the reaction,



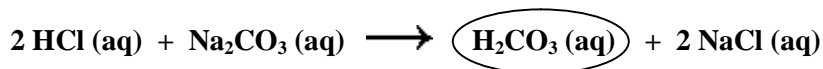
should instead be written



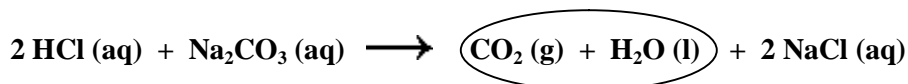
Example 3 – Formation of a Gas

When aqueous solutions of hydrochloric acid and sodium carbonate are mixed, bubbles of a gaseous product are observed.

Formula Equation:

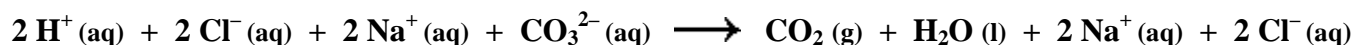


The initial product, H_2CO_3 , is a weak electrolyte and is also one of the unstable products that decompose, forming CO_2 and H_2O , so the above equation becomes,



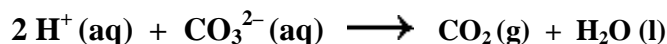
Complete Ionic Equation:

Since CO_2 and H_2O are molecular substances, we leave their formulas together. HCl is a strong electrolyte, so it is separated into its positive and negative ion in solution. The compounds Na_2CO_3 and NaCl are soluble ionic substances, so their formulas are correspondingly separated into their positive and negative ions:



Net Ionic Equation:

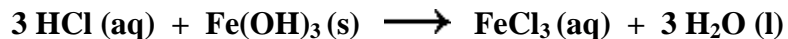
Eliminating the spectator ions (Na^+ and Cl^-) from the ionic equation gives the net ionic equation:



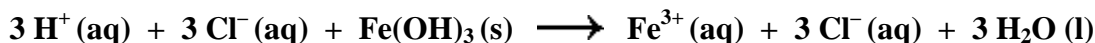
Example 4 – Reaction with a Solid Reactant

Solid hydroxide compounds, such as iron(III) hydroxide, react with hydrochloric acid.

Molecular Equation:

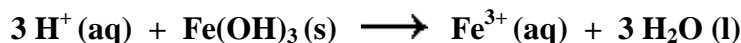


Complete Ionic Equation:



Since Fe(OH)_3 is insoluble (Rule 9), it is not separated into its positive and negative ions.

Net Ionic Equation:



EXPERIMENTAL PROCEDURE

Work individually (or in a group of 2) to carry out each of the reactions and note your observations. The reactions may be done by combining each reactant in a test tube, or you may use a smaller scale procedure, adding a few drops of each reactant to the well of a spot plate.

1. Obtain one clean spot plate (or obtain 10 clean small test tubes in a test tube rack). For each of the reactions indicated in your report sheet, add 4-5 drops of each of the two specified reactants and mix well.

If the solution turns cloudy, this is due to the formation of fine, suspended solid particles of a solid precipitate. Formation of bubbles, and a new odor, is evidence of a gaseous product.

2. Record your observations on the data sheet. If a reaction occurs, note the color of the solid formed as well as any other observations. If no visible reaction is observed, indicate so on the report sheet.
Note: Lack of a visible reaction does NOT guarantee that no reaction occurred. A soluble, colorless product such as water will not be visually apparent.

Do one reaction at a time and be very careful not to mix up droppers (if used) as this would lead to contamination of the reagent bottles. The solutions are at a concentration of "0.10 M" or similar, where "M" stands for the concentration unit molarity, which means moles of solute per liter of solution.

After completing all the reactions, but before disposing the contents of the test tubes or spot plates, discuss your observations with your lab partner and work together to write the formula equations on the report sheet.

3. For each reaction write the balanced molecular equation, complete ionic equation, and net ionic equation.

Common Ions

| Name | Formula | Charge | Name | Formula | Charge | Name | Formula | Charge |
|--------------------------------|------------------------|--------|---|------------------------------------|---------------|----------------------|------------------------------|--------|
| aluminum | Al^{3+} | +3 | magnesium | Mg^{2+} | +2 | carbonate | CO_3^{2-} | -2 |
| ammonium | NH_4^+ | +1 | manganese (II) | Mn^{2+} | +2 | chlorate | ClO_3^- | -1 |
| barium | Ba^{2+} | +2 | manganese (III) | Mn^{3+} | +3 | chloride | Cl^- | -1 |
| cadmium | Cd^{2+} | +2 | mercury (I) (mercurous) <i>{See note}</i> | Hg_2^{2+} | $2 \times +1$ | chromate | CrO_4^{2-} | -2 |
| calcium | Ca^{2+} | +2 | mercury (II) (mercuric) | Hg^{2+} | +2 | cyanate | OCN^- | -1 |
| cesium | Cs^+ | +1 | potassium | K^+ | +1 | cyanide | CN^- | -1 |
| chromium (II) | Cr^{2+} | +2 | rubidium | Rb^+ | +1 | dichromate | $\text{Cr}_2\text{O}_7^{2-}$ | -2 |
| chromium (III) | Cr^{3+} | +3 | scandium (III) | Sc^{3+} | +3 | dihydrogen phosphate | H_2PO_4^- | -1 |
| cobalt (II) (cobaltous) | Co^{2+} | +2 | silver | Ag^+ | +1 | fluoride | F^- | -1 |
| cobalt (III) (cobaltic) | Co^{3+} | +3 | sodium | Na^+ | +1 | hydroxide | OH^- | -1 |
| cobalt (IV) | Co^{4+} | +4 | tin (II) (stannous) | Sn^{2+} | +2 | iodate | IO_3^- | -1 |
| copper (I) (cuprous) | Cu^+ | +1 | tin (IV) (stannic) | Sn^{4+} | +4 | iodide | I^- | -1 |
| copper (II) (cupric) | Cu^{2+} | +2 | titanium (II) | Ti^{2+} | +2 | nitrate | NO_3^- | -1 |
| gold (I) (aurous) | Au^+ | +1 | titanium (III) | Ti^{3+} | +3 | nitrite | NO_2^- | -1 |
| gold (III) (auric) | Au^{3+} | +3 | titanium (IV) | Ti^{4+} | +4 | nitride | N^{3-} | -3 |
| hydrogen <i>{See note}</i> | H^+ | +1 | vanadium (II) | V^{2+} | +2 | oxalate | $\text{C}_2\text{O}_4^{2-}$ | -2 |
| hydronium <i>{See note}</i> | H_3O^+ | +1 | vanadium (III) | V^{3+} | +3 | oxide | O^{2-} | -2 |
| iron (II) (ferrous) | Fe^{2+} | +2 | vanadium (IV) | V^{4+} | +4 | permanganate | MnO_4^- | -1 |
| iron (III) (ferric) | Fe^{3+} | +3 | zinc | Zn^{2+} | +2 | phosphate | PO_4^{3-} | -3 |
| lead (II) (plumbous) | Pb^{2+} | +2 | acetate | $\text{C}_2\text{H}_3\text{O}_2^-$ | -1 | sulfate | SO_4^{2-} | -2 |
| lead (IV) (plumbic) | Pb^{4+} | +4 | bicarbonate (hydrogen carbonate) | HCO_3^- | -1 | sulfite | SO_3^{2-} | -2 |
| lithium | Li^+ | +1 | bisulfate (hydrogen sulfate) | HSO_4^- | -1 | sulfide | S^{2-} | -2 |
| nickel(II) (nickelous) | Ni^{2+} | +2 | bromate | BrO_3^- | -1 | thiocyanate | SCN^- | -1 |
| potassium | K^+ | +1 | bromide | Br^- | -1 | thiosulfate | $\text{S}_2\text{O}_3^{2-}$ | -2 |

A note about hydrogen and hydronium: Rarely does hydrogen ion exist on its own. When H^+ is written in equations or textbooks, it usually is a simplified way of saying H_3O^+ . Water, H_2O , is constantly breaking up to form trace amounts of hydroxide (OH^-) and hydronium (H_3O^+) ions.

A note about mercury: Mercury (I) ion exists as a diatomic unit.

EXPERIMENT 5 – Double Replacement Reactions

REPORT FORM

Name _____

Instructor _____

Date _____

1. barium chloride (aq) + sodium sulfate (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

2. silver nitrate (aq) + sodium chloride (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

3. lead(II) nitrate (aq) + potassium iodide (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

4. sodium carbonate (aq) + hydrochloric acid (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

5. sodium hydroxide (aq) + hydrochloric acid (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

6. barium chloride (aq) + sodium carbonate (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

7. sodium hydroxide (aq) + acetic acid (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

8. sodium hydroxide (aq) + ammonium chloride (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

9. lead(II) nitrate (aq) + sodium sulfide (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

10. sodium sulfide (aq) + hydrochloric acid (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

11. copper(II) sulfate (aq) + sodium hydroxide (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

12. lead (II) nitrate (aq) + potassium chromate (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

13. iron(III) nitrate (aq) + sodium hydroxide (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

14. silver nitrate (aq) + sodium hydroxide (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

15. sodium chloride (aq) + calcium nitrate (aq)

Observation:

Balanced Formula Equation:

Complete Ionic Equation:

Net Ionic Equation:

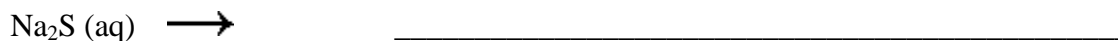
EXPERIMENT 5

Name: _____

Pre-Laboratory Questions and Exercises

Due before lab begins. Answer in the space provided.

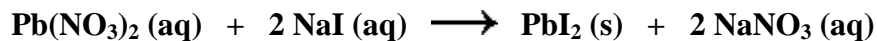
1. Write the positive and negative ions that result when the following compounds are dissolved in aqueous solution:



2. Using the solubility rules, predict the solubility of each of the following compounds in water. (S = soluble, I = insoluble)

a) CaCO_3 _____ b) $\text{Al}(\text{OH})_3$ _____ c) Hg_2Cl_2 _____ d) $\text{Cu}(\text{NO}_3)_2$ _____

3. Write the complete ionic and net ionic equations for the following reaction:



Complete Ionic Equation:

Net Ionic Equation: