EXPERIMENT **19**

Net Ionic Equations: A Study Assignment

Performance Goals

19–1 Distinguish between a conventional equation, an ionic equation, and a net ionic equation.

19–2 Identify spectators in an ionic equation.

19–3 Given information from which you may write the conventional equation for one of the following types of reactions in aqueous solution, write the net ionic equation:

  a) Ion combination reactions that produce a precipitate;

  b) Ion combination reactions that produce a molecular product;

  c) Reactions that yield a gas;

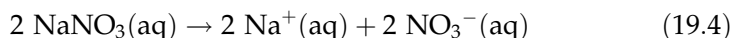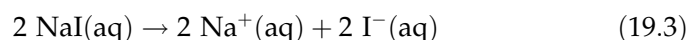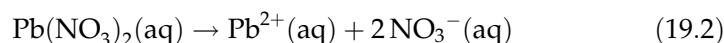 **d)** Oxidation–reduction reactions that may be described by a ''single replacement'' equation.

CHEMICAL OVERVIEW

The conventional equation by which a chemical change is described is adequate for most purposes, but for reactions of ionic compounds in aqueous (water) solution, the typical molecular equation has serious shortcomings. Specifically, a conventional equation may indicate formulas of reactants and products that are not present, and omit completely the formulas of the ions that are the real reactants and products. If the substances in the conventional equation that are actually present as dissociated ions are written in the form of their ions, the result is an **ionic equation.** Frequently the same ions appear on both sides of an ionic equation. Though they are present at the scene of a chemical reaction, these ions experience no chemical change themselves. They are called **spectator ions,** or simply **spectators.** Eliminating the spectators from an ionic equation leaves the **net ionic equation,** an equation that includes only the substances that actually participate in the reaction.

To illustrate these equations, consider the reaction between solutions of lead(II) nitrate and sodium iodide. When the two clear solutions are combined, insoluble lead(II) iodide precipitates. The conventional equation is

$$Pb(NO_3)_2(aq) + 2\,NaI(aq) \rightarrow PbI_2(s) + 2\,NaNO_3 \tag{19.1}$$
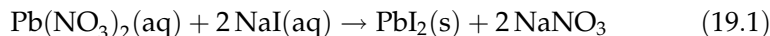
Three of the compounds in the equation are soluble ionic salts that dissociate into ions when dissolved. In other words, no real solute particles having the formulas $Pb(NO_3)_2$, NaI, or $NaNO_3$ are actually present in the reaction vessel. The species actually present are the ions resulting from those species. These ions, both their identity and number, can be found by writing the equation for the dissociation of that number of solute species appearing in the conventional equation:

$$Pb(NO_3)_2(aq) \rightarrow Pb^{2+}(aq) + 2\,NO_3^-(aq) \tag{19.2}$$

$$2\,NaI(aq) \rightarrow 2\,Na^+(aq) + 2\,I^-(aq) \tag{19.3}$$

$$2\,NaNO_3(aq) \rightarrow 2\,Na^+(aq) + 2\,NO_3^-(aq) \tag{19.4}$$

Replacing the compound formulas of dissolved substances in the conventional equation with the ions really present gives the ionic equation:

$$Pb^{2+}(aq) + 2\,NO_3^-(aq) + 2\,Na^+(aq) + 2\,I^-(aq) \rightarrow PbI_2(s) \tag{19.5}$$
$$+ 2\,Na^+(aq) + 2\,NO_3^-(aq)$$
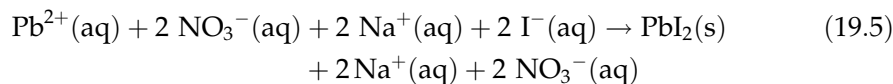
The ionic equation contains two sodium ions and two nitrate ions on both sides. These are the spectators that do not undergo any change, even though they are in the solution where chemical change is occurring. Eliminating the spectators from the ionic equation gives the net ionic equation:

$$Pb^{2+}(aq) + 2\,I^-(aq) \rightarrow PbI_2(s) \tag{19.6}$$
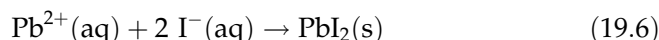
The net ionic equation isolates the two reactants that actually change chemically and the single new substance produced.

The above example does more than simply illustrate a net ionic equation. It also shows how that equation is developed. There are three steps in the procedure:

1. Write the conventional equation, including designations of state or solutions, (s) for solid, (ℓ) for liquid, (g) for gas, and (aq) for dissolved species (Equation 19.1). Be sure the equation is balanced.*

2. Write the ionic equation, replacing those species that ionize in aqueous solution with ions actually present (Equation 19.5). Check to be sure the equation is still balanced, both in atoms and in electrical charge.

3. Write the net ionic equation by eliminating the spectators (Equation 19.6). If necessary, reduce coefficients to their lowest integral (whole number)

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*If the purpose is simply to write the net ionic equation for a reaction, it is not essential that the conventional and ionic equations be balanced. Most instructors, however, require balancing all three equations. That procedure will be followed in this book; you should follow the procedure recommended by your instructor.

values. Check to be sure the equation is still balanced, both in atoms and electrical charge.

Regardless of what you may or may not have been doing up to this point in your chemistry course, it is particularly important that you include designations of state or solution in your conventional equation. These designations identify the formulas that *may* be rewritten in the form of dissociated ions in the ionic equation. *Under no circumstances do you change the formula of a solid, liquid, or gas; those substances appear in the ionic and net ionic equations in exactly the same form as they appear in the conventional equation.* It follows that you must have sufficient information in the description of the reaction to determine the proper designations for all substances. Reactant designations are determined by careful reading of the description of the reaction. Product designations may or may not be indicated. It is generally left to you to decide whether a product formed by the combination of ions is or is not soluble, and, even if it is soluble, whether it does or does not dissociate into ions.

Solubility. The solubility of an ion combination product may be determined by referring to Table 19.1. Directions for its use appear beneath the table. Notice that the intersection of the horizontal lead(II) ion line with the vertical iodide ion column contains the letter "s," indicating that lead(II) iodide is an insoluble solid, written as $PbI_2(s)$ in an equation. The intersection of the horizontal sodium ion line with the vertical nitrate ion column shows "aq," indicating that sodium nitrate is soluble in water, and is written $NaNO_3(aq)$ in an equation. It is worth noting that all entries in the sodium ion line and the nitrate ion column are designated "aq," showing that all sodium salts and all nitrate salts included in the table are soluble. These and several other useful generalizations are listed in the following "solubility rules":

1. All ammonium and alkali metal salts are soluble.
2. All nitrates are soluble.
3. All chlorides and bromides are soluble except those of Ag^+, Hg_2^{2+}, and Pb^{2+} ions.
4. All sulfates are soluble except those of Ca^{2+}, Sr^{2+}, Ba^{2+}, Hg_2^{2+}, Hg^{2+}, Pb^{2+}, and Ag^+ ions.
5. All carbonates are insoluble except ammonium carbonate and the carbonates of the alkali metals.
6. All hydroxides are insoluble except those of the alkali metals, $Ba(OH)_2$ and $Sr(OH)_2$.
7. All sulfides are insoluble except those of the alkali metals and ammonium sulfide.

Molecular Products—Weak Acids. Sometimes the hydrogen ion from a strong acid will combine with the anion of a second reactant to form a molecular product. If the anion is the hydroxide ion, the molecular product is water, $H_2O(\ell)$, which appears in that form in all conventional and ionic equations. Another molecular product forms when the hydrogen ion from a strong acid combines with the anion of a weak acid. These acids ionize

Table 19.1 *Solubilities of Ionic Compounds**

Cations / Anions	Acetate	Bromide	Carbonate	Chlorate	Chloride	Fluoride	Hydrogen Carbonate	Hydroxide	Iodide	Nitrate	Nitrite	Phosphate	Sulfate	Sulfide	Sulfite
Aluminum	S	aq		aq	aq	S		S	—	aq		S	aq	—	
Ammonium	aq	aq	aq	aq	aq	aq	aq	—	aq	aq	aq	aq	aq	aq	aq
Barium	—	aq	S	aq	aq	S		aq	aq	aq	aq	S	S	—	S
Calcium	aq	aq	S	aq	aq	S		S	aq	aq	aq	S	S	S	S
Cobalt(II)	aq	aq	S	aq	aq	—		S	aq	aq		S	aq	S	S
Copper(II)	aq	—			aq	aq		S		aq		S	aq	S	
Iron(II)	aq	aq	S		aq	S		S	aq	aq		S	aq	S	S
Iron(III)	—	aq			aq	S		S	—	aq		S	aq	—	
ead(II)	aq	S	S	aq	S	S		S	S	aq	aq	S	S	S	S
ithium	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	S	aq	aq	
agnesium	aq	aq	S	aq	aq	S		S	aq	aq	aq	S	aq	—	aq
ickel		aq	S	aq	aq	aq		S	aq	aq		S	aq	S	S
otassium	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq
Silver	S	S	S	aq	S	aq		—	S	aq	S	S	S	S	S
Sodium	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq
inc	aq	aq	S	aq	aq	aq		S	aq	aq		S	aq	S	S

*To determine the solubility of an ionic compound, locate the intersection of the horizontal row for the cation and the vertical column for the anion. An "aq" in that box indicates that the compound is soluble in water to a molarity of 0.1 or more at 20°C. An "s" indicates that the compound is not soluble to that concentration, but remains in the solid state or precipitates if the ions are combined. A blank space indicates lack of data, and a dash (—) identifies an unstable substance.

only slightly, and when writing ionic equations are considered to exist in molecular form. Hydrofluoric acid, HF, for example, happens to be an extremely corrosive acid, but it is also a weak acid because the individual solute particles in its water solution are predominantly un-ionized HF molecules. Its proper designation is therefore HF(aq) in ionic equations as well as conventional equations. In other words, *a weak acid is not separated into ions in writing an ionic equation, but is written in molecular form.*

All of this makes it necessary for you to distinguish between weak acids and strong acids. Fortunately, there are only a few strong acids—few enough for you to memorize them easily. They are the three most common acids, sulfuric (H_2SO_4), nitric (HNO_3), and hydrochloric (HCl), and three other acids containing halogens, hydrobromic (HBr), hydroiodic (HI), and perchloric ($HClO_4$). If any of these six acids appears in a conventional equation, it must be broken down to ions in an ionic equation. All other

acids are written in molecular form in both conventional and ionic equations.*

Unstable Ion Combination Products.   There are three common substances that are unstable in water solution and must be written in both the conventional and ionic equations in the form of their final decomposition products. These are carbonic acid, H_2CO_3, sulfurous acid, H_2SO_3, and what is commonly called ''ammonium hydroxide,'' whose formula is usually written NH_4OH. Carbonic and sulfurous acids break down into water plus carbon dioxide or sulfur dioxide, respectively. Carbon dioxide is relatively insoluble and bubbles out of the solution as a gas, $CO_2(g)$. Sulfur dioxide is relatively soluble and is written as $SO_2(aq)$ in the equation. ''Ammonium hydroxide,'' NH_4OH, never really forms from NH_4^+ and OH^- ions; instead, an equilibrium is reached, $NH_4^+(aq) + OH^-(aq) \rightleftarrows NH_3(aq) + H_2O(\ell)$, in which the products are by far the predominant species and are written in molecular form in all equations.

    To summarize:

What breaks down (forms ions)?

1. Strong acids;
2. Strong bases;
3. Soluble salts.

What does not break down?

1. Weak acids (remember, if it is not one of the strong acids, it is considered a weak acid!);
2. Weak bases;
3. Gases;
4. Solids;
5. Elements.

EXAMPLES

The following examples are in the form of a program in which you learn by answering a series of questions. Obtain an opaque shield—a piece of cardboard or a folded piece of paper you cannot see through—that is wide enough to cover this page. In each example, place the shield on the book page so it covers everything beneath the first dotted line that runs across the page. Read to that point, and write in the space provided whatever is asked. Then lower the shield to the next dotted line. The material exposed will begin with the correct response to the question you have just answered. Compare this answer to yours, looking back to correct any misunderstanding if the two are different. When you fully understand the first step, read to the next dotted line and proceed as before.

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*Some fairly common acids, notably phosphoric, oxalic, and chloric acids, are ''borderline'' acids that might be considered either strong or weak, depending on some arbitrary classification standard. We will avoid these acids in this exercise.

### Example 1

Write the net ionic equation for the reaction between solutions of barium bromide and ammonium sulfate.

Begin with the conventional equation. You will have to predict the products of the reaction, which is not difficult if you recall that the reaction between two ionic substances is frequently described by a ''double displacement'' type of equation, in which the positive ion of one reactant combines with the negative ion of the other, and vice versa—a ''change partners'' type of equation. Be sure to include state or solution designations, using Table 19.1 to determine the proper designations for the products.

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1a. $BaBr_2(aq) + (NH_4)_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ NH_4Br(aq)$
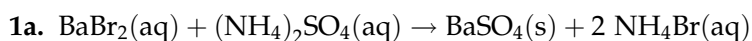
The (aq) designations for the reactants are indicated by the word ''solutions'' in the statement of the question. The intersection of the barium ion line and sulfate column line in Table 19.1 has an ''s'' in it, indicating that barium sulfate is a solid (precipitate), and therefore its designation in the equation is (s). The ammonium ion and bromide ion intersection shows ''aq,'' indicating that the compound is soluble in water.

Now write the ionic equation by replacing those species that are in solution, designated (aq), by the ionic particles actually present in the solution. Remember, only those substances designated (aq) may be changed in this step. Be sure your equation is balanced.

....

1b. $Ba^{2+}(aq) + 2\ Br^-(aq) + 2\ NH_4{}^+(aq) + SO_4{}^{2-}(aq)$
$$\rightarrow BaSO_4(s) + 2\ Br^-(aq) + 2\ NH_4{}^+(aq)$$
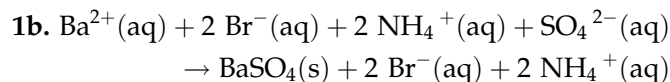
The $BaSO_4(s)$, a solid, remains in the same form in the ionic equation as it appeared in the conventional equation. The other substances, all ionic compounds, dissociate into their ions.

From here all you must do is cross out the spectators and you have the net ionic equation.

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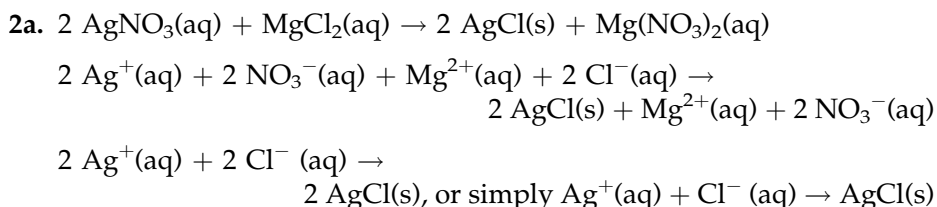**1c.** $Ba^{2+}(aq) + SO_4{}^{2-}(aq) \rightarrow BaSO_4(s)$
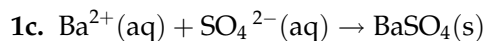
A final check on the equation shows one barium ion on each side, one sulfate ion on each side, and a total charge of zero on both sides. The importance of the charge balance will appear shortly.

The next example is similar, but slightly different. . . .

Example 2

Write all equations that result in the net ionic equation for the reaction between solutions of silver nitrate and magnesium chloride.

You know the procedure for all three steps. Proceed, but be careful on the last step.... .

....

2a. $2\ AgNO_3(aq) + MgCl_2(aq) \rightarrow 2\ AgCl(s) + Mg(NO_3)_2(aq)$

$2\ Ag^+(aq) + 2\ NO_3^-(aq) + Mg^{2+}(aq) + 2\ Cl^-(aq) \rightarrow$
$$2\ AgCl(s) + Mg^{2+}(aq) + 2\ NO_3^-(aq)$$

$2\ Ag^+(aq) + 2\ Cl^-(aq) \rightarrow$
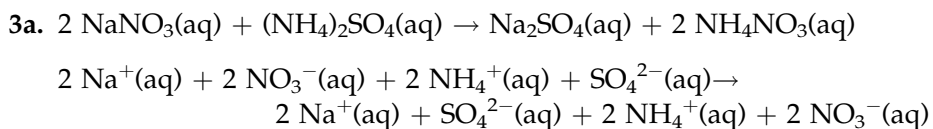$$2\ AgCl(s), \text{ or simply } Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

The net ionic equation that results from the elimination of the magnesium and nitrate ion spectators has coefficients of 2 for all species. Step 3 in the procedure for writing net ionic equations (page 246) states that all coefficients should be expressed in their lowest terms. Dividing the first equation in the last line above by 2 yields the second equation.

At some point you should become aware of the possible combination of solutions that leads to the conditions in the next example.

Example 3

Write the net ionic equation for the reaction, if any, that occurs when a solution of sodium nitrate is poured into a solution of ammonium sulfate.

Proceed, but when you come to something you haven't seen before, think about it....

....

3a. $2\ NaNO_3(aq) + (NH_4)_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\ NH_4NO_3(aq)$

$2\ Na^+(aq) + 2\ NO_3^-(aq) + 2\ NH_4^+(aq) + SO_4^{2-}(aq) \rightarrow$
$$2\ Na^+(aq) + SO_4^{2-}(aq) + 2\ NH_4^+(aq) + 2\ NO_3^-(aq)$$

There is no net ionic equation because every species in the ionic equation is a spectator. In other words, there is no reaction, no chemical change, no reactant destroyed, no new substance formed. The final mixture contains the four ions that were initially divided between two solutions. When you encounter a question such as this, simply state ''No reaction.''

A neutralization reaction is one between an acid and a base, yielding water and a salt as its products. The next example gives the net ionic equation that describes all neutralizations between solutions of a strong acid and a strong base.

Example 4

Write the net ionic equation for the reaction between aqueous potassium hydroxide and sulfuric acid.

The procedure is the same as before, but there will be a small difference that shows up in your state designations. Complete the equation.
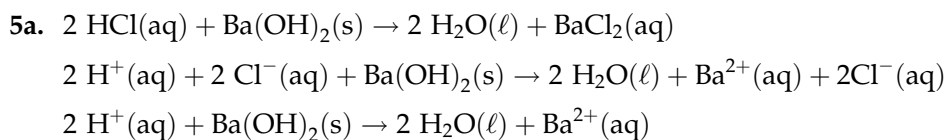
....

4a. $H_2SO_4(aq) + 2\ KOH(aq) \rightarrow 2\ H_2O(\ell) + K_2SO_4(aq)$

$2\ H^+(aq) + SO_4^{2-}(aq) + 2\ K^+(aq) + 2\ OH^-(aq)$
$$\rightarrow 2\ H_2O(\ell) + SO_4^{2-}(aq) + 2\ K^+(aq)$$

$2\ H^+(aq) + 2\ OH^-(aq) \rightarrow 2\ H_2O(\ell)$, or simply $H^+(aq) + OH^-(aq)$
$$\rightarrow H_2O(\ell)$$
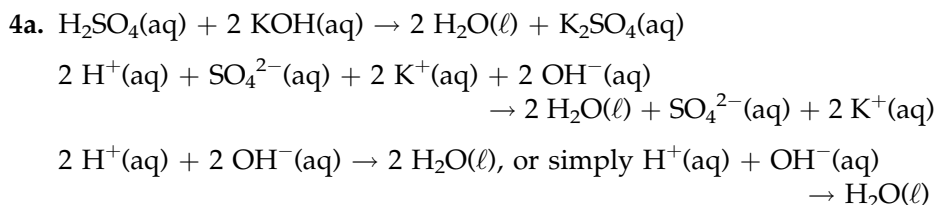
The difference in state designation is that water is a liquid, designated (ℓ), rather than a solid precipitate as in Examples 1 and 2.

A common error frequently appears in writing the net ionic equation for a reaction involving sulfuric acid, as in Example 4. In writing the ionic equation, the acid is incorrectly broken into the sulfate ion and an $H_2^+(aq)$ ion, which does not exist. The formula for an aqueous hydrogen ion is $H^+(aq)$, whether it comes from H_2SO_4, HCl, or any other acid. The fact that elemental hydrogen is diatomic, H_2, has nothing to do with the aqueous hydrogen ion.

Example 5

Write the net ionic equation for the reaction between hydrochloric acid and solid barium hydroxide.

The reaction is very similar to Example 4, but you must read carefully....

....

5a. $2\ HCl(aq) + Ba(OH)_2(s) \rightarrow 2\ H_2O(\ell) + BaCl_2(aq)$

$2\ H^+(aq) + 2\ Cl^-(aq) + Ba(OH)_2(s) \rightarrow 2\ H_2O(\ell) + Ba^{2+}(aq) + 2Cl^-(aq)$

$2\ H^+(aq) + Ba(OH)_2(s) \rightarrow 2\ H_2O(\ell) + Ba^{2+}(aq)$
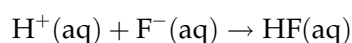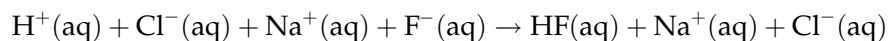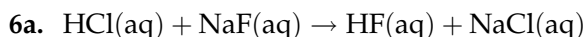
In this example it is stated that *solid* barium hydroxide is the reactant, even though barium hydroxide is sufficiently soluble to be classified (aq) in Table 19.1. Watch for words in the statement of a reaction that indicate the state of a reactant, and be sure to use that state in the equation.

The next neutralization reaction yields a slightly different net ionic equation.

Example 6

Write the net ionic equation for the reaction between hydrochloric acid and a solution of sodium fluoride.

Proceed as before, but remember the discussion of weak acids on page 247.

....

6a. $HCl(aq) + NaF(aq) \rightarrow HF(aq) + NaCl(aq)$

$H^+(aq) + Cl^-(aq) + Na^+(aq) + F^-(aq) \rightarrow HF(aq) + Na^+(aq) + Cl^-(aq)$
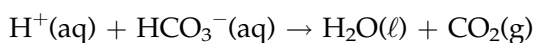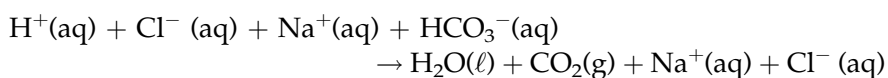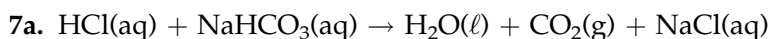
$H^+(aq) + F^-(aq) \rightarrow HF(aq)$

In this case the product HF(aq) is a weak acid—recognized as such because it is not one of the six strong acids—and therefore does not separate into ions.

The base used to neutralize an acid is not necessarily a hydroxide, as the next example shows.

Example 7

Write the net ionic equation for the reaction between hydrochloric acid and a solution of sodium hydrogen carbonate.

Again proceed, but this time remember the discussion of unstable ion combination products on page 249.

....

7a. $HCl(aq) + NaHCO_3(aq) \rightarrow H_2O(\ell) + CO_2(g) + NaCl(aq)$

$H^+(aq) + Cl^-(aq) + Na^+(aq) + HCO_3^-(aq)$
$$\rightarrow H_2O(\ell) + CO_2(g) + Na^+(aq) + Cl^-(aq)$$

$H^+(aq) + HCO_3^-(aq) \rightarrow H_2O(\ell) + CO_2(g)$

This time the ion combination product is unstable carbonic acid, H_2CO_3, which decomposes into water and carbon dioxide.

The reaction between an active metal and an acid releases hydrogen gas and leaves a solution of a salt behind. This kind of oxidation–reduction reaction is described by a conventional equation that is sometimes called a single replacement equation, in which one element appears to replace another.

Example 8

Write the net ionic equation for the reaction between zinc and sulfuric acid.

....

8a. $Zn(s) + H_2SO_4 \ (aq) \rightarrow H_2(g) + ZnSO_4(aq)$

$Zn(s) + 2H^+(aq) + SO_4{}^{2-} \ (aq) \rightarrow H_2(g) + Zn^{2+}(aq) + SO_4{}^{2-}(aq)$

$Zn(s) + 2 \ H^+(aq) \rightarrow H_2(g) + Zn^{2+}(aq)$
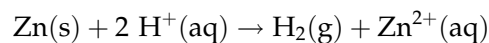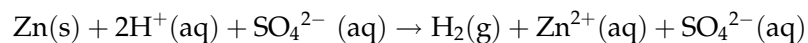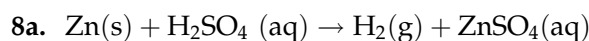
In the conventional equation it looks as if zinc replaces hydrogen to form a compound with the sulfate ion. The net ionic equation shows the true character of the reaction, in which two electrons move from a zinc atom to two hydrogen ions. It is this transfer of electrons that classifies the reaction as an oxidation–reduction reaction.

Example 8 differs from all of the earlier examples in that each side of the equation has a net electrical charge, instead of having all charges cancel out to zero. The requirement is tha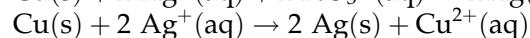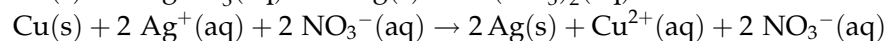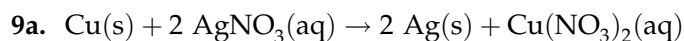t the charges be *balanced*, but not necessarily at zero. If you balance the conventional equation and correctly represent the number of ions in the ionic equation that come from the conventional equation, the charges in the net ionic equation will already be balanced. They should be checked, however, and not assumed. If the charge is not balanced, something is wrong someplace in the three steps. It is possible to write a net ionic equation that is balanced in atoms but not in charge. The final example is such a reaction.

Example 9

Write the net ionic equation for the reaction between copper and a solution of silver nitrate.

The conventional equation for this oxidation–reduction reaction between copper and aqueous silver nitrate is a single replacement equation in which the copper(II) ion is produced. This hint should enable you to write the three equations.

....

9a. $Cu(s) + 2 \ AgNO_3(aq) \rightarrow 2 \ Ag(s) + Cu(NO_3)_2(aq)$

$Cu(s) + 2 \ Ag^+(aq) + 2 \ NO_3{}^-(aq) \rightarrow 2 \ Ag(s) + Cu^{2+}(aq) + 2 \ NO_3{}^-(aq)$

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

An error in the first or second step could have produced $Cu(s) + Ag^+(aq) \rightarrow Ag(s) + Cu^{2+}(aq)$ as a net ionic equation that is balanced in atoms. The net charge on the left side of the equation is $1+$, however, and the charge on the right side is $2+$. The equation is therefore not balanced, even though both copper and silver check out.

DEMONSTRATION (OPTIONAL)

Observe the following process carried out by the instructor, and write the net ionic equation for the reaction. Also, explain the reason for the changes during the titration.

Procedure.   Pour 30 to 35 mL of clear, saturated barium hydroxide solution (carefully decanted from a bottle containing excess solute on the bottom) into a 400-mL beaker. Add about 100 mL of deionized water and 3 drops of phenolphthalein indicator. Immerse the electrodes of a conductivity-sensing device (see Figure 18.1) into the solution.

In a solution, the current-carrying species are ions; hence, good conduction indicates a great number of ions. Conversely, poor conduction shows that a relatively small number of ions are present. Observe whether the barium hydroxide solution in the beaker is a good or a poor conductor of electricity.

From a buret, slowly add a 1 M sulfuric acid solution to the contents of the beaker. Stir the solution continuously and observe any changes that occur.

Name _____   *Date* _____   *Section* _____

Experiment 19

Advance Study Assignment

1. An ionic equation has $2H^+(aq)$ on the left side and $H_2(g)$ on the right. Are these species spectators? Explain.

2. Which of the following substances are *not* ionized in an aqueous solution? Explain.

   a) HI;   b) KBr;   c) H_2SO_4;   d) HNO_2;   e) AgCl;   f) $Ca(NO_3)_2$

3. Consider the statement: After balancing a net ionic equation, the sum of the charges on each side of the equation should be equal. Is this true or false? Give an example.

Name   Date   Section

Experiment 19

Work Page

Instructions: For each reaction described below, write the conventional equation, ionic equation, and net ionic equation. Include designations of state or solution in each equation.

1. Lead(II) nitrate and magnesium sulfate solutions are combined.

2. Barium metal is dropped into hydrochloric acid.

3. Potassium hydroxide solution reacts with nitric acid.

4. Zinc chloride solution is poured into a solution of ammonium carbonate.

5. Ammonium chloride and potassium hydroxide solutions are combined.

6. Magnesium chloride solution is mixed with nickel nitrate solution.

7. Sulfuric acid reacts with a solution of magnesium acetate, $Mg(C_2H_3O_2)_2$.

8. Zinc reacts with a solution of nickel sulfate.

9. Cobalt(II) sulfate and lithium sulfide solutions are combined.

10. Potassium nitrite solution is added to sulfuric acid.

11. Solid nickel carbonate is dropped into nitric acid.

12. Iron(III) chloride and aluminum sulfate solutions are mixed.

13. Hydrochloric acid is poured into a solution of lithium carbonate.

Name   _Date_   _Section_

Experiment 19

Work Page

14. Hydrogen is released when sodium reacts with water.

15. Hydroiodic acid reacts with a solution of ammonium sulfite.

16. Solid copper(II) hydroxide is ''dissolved'' by hydrochloric acid.

17. Sodium hydroxide solution is poured into a solution of cobalt(II) chloride.

18. Calcium metal reacts with a solution of iron(II) bromide.

19. Hydrochloric acid reacts with a solution of sodium acetate, $NaC_2H_3O_2$.

20. Aluminum bromide and ammonium fluoride solutions are combined.

Demonstration Questions (Optional)

1. What were the color and conductivity of the initial barium hydroxide and phenolphthalein solution?

2. What changes occurred during the titration?

3. Why did the pink color disappear?

4. Why did the light go out?

5. Why did the light come back on?

6. Write the conventional and the net ionic equations for the reaction.

Name   _Date_   _Section_

Experiment 19

Report Sheet

Instructions: For each reaction described below, write the conventional equation, ionic equation, and net ionic equation. Include designations of state or solution in each equation.

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8. Zinc reacts with a solution of nickel sulfate.

9. Cobalt(II) sulfate and lithium sulfide solutions are combined.

10. Potassium nitrite solution is added to sulfuric acid.

11. Solid nickel carbonate is dropped into nitric acid.

12. Iron(III) chloride and aluminum sulfate solutions are mixed.

13. Hydrochloric acid is poured into a solution of lithium carbonate.

Name   *Date*   *Section*

Experiment 19

Report Sheet

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