

# *Conductivity and Net Ionic Equations*

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## *Objective*

To use electrical conductivity as a way of determining the number of free ions present in a substance and to use this information to draw conclusions regarding the type of bonding present in the substance.

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## *Introduction*

### **Types of Bonding**

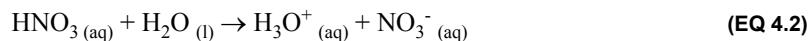
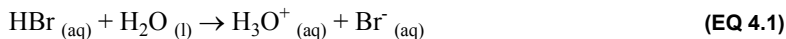
There are three basic types of bonds. They are ionic, covalent, and polar covalent, which is a hybrid of an ionic and a covalent bond. Ionic bonds are formed between elements with very different electronegativities. **Electronegativity** is the ability of an atom to draw electrons towards itself in a chemical bond. Generally, ionic compounds form between metals and non-metals and are identified by the transfer of an electron from the metal to the non-metal to form charged ions which are held together by electrostatic interactions known as ionic bonds. Covalent bonds form between elements with similar electronegativities. Generally, covalent bonds form between non-metals and are characterized by the sharing of electron pairs between the atoms. Polar bonds are formed between two elements with different electronegativities, but which still share the electrons albeit unevenly.

### **Electrolytes versus Non-electrolytes**

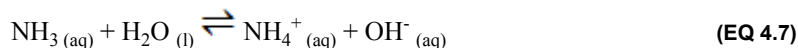
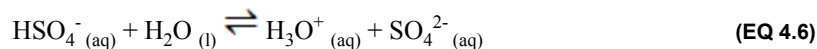
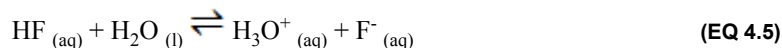
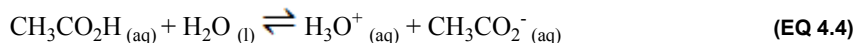
Substances may be classified by their electrical conductivity. **Electronic conduction** is a type of electrical conductivity that occurs in metals where charge is carried by electrons. In **ionic conduction** the charge is carried by ions. Substances which can conduct electricity are called **electrolytes**. When ionic compounds are melted or dissolved in water, they form mobile ions that are able to conduct electricity. Polar covalent compounds such as acids and bases will sometimes dissociate or break apart in aqueous solution to form ions as well. There are two types strong electrolyte and weak electrolytes.

Substances which do not conduct electricity are called **non-electrolytes**. Examples of non-electrolytes are covalent molecules such as sucrose or table sugar ( $C_{12}H_{22}O_{11}$ ) and acetone ( $CH_3CO-CH_3$ ). These substances are non-electrolytes because they are not composed of ions and cannot conduct an electric current. Ionic compounds in their crystalline form are also considered to be non-electrolytes because even though they are composed of ions, the ions are not able to move freely through the crystal and therefore are not able to conduct electricity.

**Strong electrolytes** dissociate ~100% in aqueous solution. In chemical equations where ions are included (total and net ionic equations) write the ionized components, since they are the major species present in the solution. Examples of these are shown below:



**Weak electrolytes** are substances that are able to conduct electricity, but conduct poorly. Examples of weak electrolytes are molecular substances that dissociate to a small extent such as weak acids and bases, and ionic compounds that have limited water solubility. They are **slightly ionizable substances**. In chemical equations where ions are included (total and net ionic equations) write the complete weak acid formula, not the ions. The ions are only a minor component in the total solution. Examples are shown below:



As you complete the experiment, notice the difference in conductivity between strong and weak acids and bases. What does this tell you about the relative degree of dissociation? What are the principal species present in the solution? How do we write these substances in ionic equations?

### Writing Chemical Reactions

Reaction can be classified into two basic types: oxidation-reduction and double displacement.

1. Oxidation-Reduction (Redox) Reactions: Electrons are transferred from one reactant to another. Most simply oxidation is the loss of electrons and reduction is the gain of electrons.

a. Combination reactions where reactants combine to form a new substance:



b. Decomposition reactions where reactants break apart into new substances:

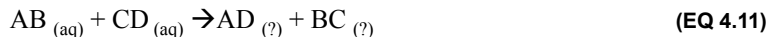


c. Single Replacement Reactions involve the reactivity of an element is related to its tendency to lose or gain electrons; that is, to be oxidized or reduced. Generally speaking:



where A is the more active element and replaces B in the compound.

2. Double Displacement Reactions (aka Ion Exchange Reactions) two aqueous solutions are mixed together to produce a precipitate, slightly ionizable substance, or a gas.



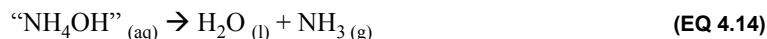
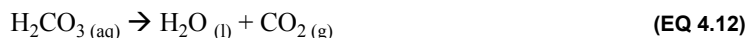
- a. Precipitate — formation of an insoluble compound. Solubility rules are given below:

#### FIGURE 4.1 Solubility Rules

- Alkali metals and  $\text{NH}_4^+$  compounds are soluble (i.e. aqueous, (aq)), this means they completely dissociate in water.
- Nitrates ( $\text{NO}_3^-$ ), acetates ( $\text{CH}_3\text{CO}_2^-$ ) (except for acetates of  $\text{Al}^{3+}$ , and  $\text{Ag}^+$ ), chlorates ( $\text{ClO}_3^-$ ), perchlorates ( $\text{ClO}_4^-$ ), and sulfates ( $\text{SO}_4^{2-}$ ) are generally soluble (except for  $\text{Sr}^{2+}$ ,  $\text{Ca}^{2+}$ ,  $\text{Ba}^{2+}$ ,  $\text{Pb}^{2+}$ , and  $\text{Hg}_2^{2+}$  sulfates).
- Chlorides ( $\text{Cl}^-$ ), bromides ( $\text{Br}^-$ ), and iodides ( $\text{I}^-$ ) are soluble (except for silver ( $\text{Ag}^+$ ), mercury(I) ( $\text{Hg}_2^{2+}$ ), and lead(II) ( $\text{Pb}^{2+}$ ) halides).
- Compounds containing the fluoride ion,  $\text{F}^-$ , are soluble except for group IIA cations and  $\text{Pb}^{2+}$ . These cations will precipitate out as a fluoride salt.
- Most compounds not included above are not soluble (i.e. precipitates/solids, (s)): i.e. sulfides ( $\text{S}^{2-}$ ), carbonates ( $\text{CO}_3^{2-}$ ), phosphates ( $\text{PO}_4^{3-}$ ), chromates ( $\text{CrO}_4^{2-}$ ), oxides ( $\text{O}^{2-}$ ), and hydroxides ( $\text{OH}^-$ ) ( $\text{Ca}(\text{OH})_2$ ,  $\text{CaO}$ ,  $\text{Sr}(\text{OH})_2$ ,  $\text{SrO}$ ,  $\text{Ba}(\text{OH})_2$ , and  $\text{BaO}$  are slightly soluble meaning that they dissociate some, but not completely in water.).

- b. Gas — bubbles or effervescence. Common gases include hydrogen gas,  $\text{H}_2_{(g)}$ , oxygen gas,  $\text{O}_2_{(g)}$ , hydrogen sulfide,  $\text{H}_2\text{S}_{(g)}$ , ammonia,  $\text{NH}_3_{(g)}$ , carbon dioxide,  $\text{CO}_2_{(g)}$ , and sulfur dioxide,  $\text{SO}_2_{(g)}$ .

When these compounds are “formed” they are unstable and decompose into gases and water:



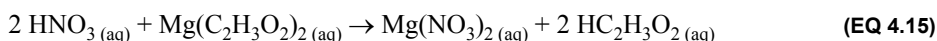
- c. Slightly ionized substance — heat usually accompanies the formation of water,  $\text{H}_2\text{O}$ , acetic acid,  $\text{HC}_2\text{H}_3\text{O}_2$ , or any other slightly ionized compound (weak acids, weak bases).

#### Writing Ionic Equations

When you write ionic equations, you need to show the principal species present in the solutions. By testing the conductivity of a variety of solutions you can determine whether the principal species are ions or undissociated or undissolved particles. For each of the substances test for electrical conductivity and determine the principal and minor species present in the solution.

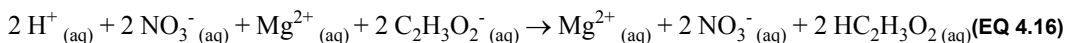
**EXAMPLE 4.1** Write the balanced conventional equation for the reaction of nitric acid and magnesium acetate.

Remember that the conventional equation shows all species as neutral compounds.



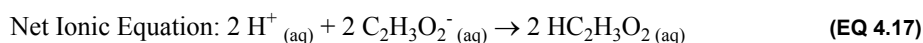
**EXAMPLE 4.2** Write the total ionic equation for the reaction.

Remember that the total ionic equation shows all species as they appear in solution.

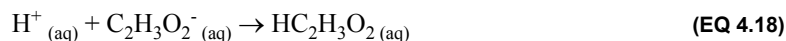


**EXAMPLE 4.3** Write the net ionic equation for the reaction.

The net ionic equation shows only the species that undergo reaction. No spectator ions.



Remember to simplify the coefficients when necessary.



**EXAMPLE 4.4** Complete the following table:

TABLE 4.5

	Conductivity	Ions (None, Few, Many)	Major species present in solution	Minor species in solution
Before Rxn: $\text{HNO}_3(\text{aq})$	Good	Many	$\text{H}^+(\text{aq}), \text{NO}_3^-(\text{aq})$	n/a
Before Rxn: $\text{Mg}(\text{C}_2\text{H}_3\text{O}_2)_2(\text{aq})$	Good	Many	$\text{Mg}^{2+}(\text{aq}), \text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$	n/a
After Reaction	Good	Many	$\text{Mg}^{2+}(\text{aq}), \text{NO}_3^-(\text{aq}), \text{HC}_2\text{H}_3\text{O}_2(\text{aq})$	$\text{H}^+(\text{aq}), \text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$

Note that water is omitted as a major species since it is a solvent.

Notice that the conductivity is good before and after the reaction due to the presence of ions in the solutions before and after the reaction.

### Information to Include in the Introduction

- Name the three principal bond types.
- Define the following terms and tell what kind of bonding is expected in solutes that are:
  - Non-electrolyte
  - Strong electrolyte
  - Weak electrolyte
- What is the meaning of the term “hydration” when used to describe what happens to an electrolyte which is dissolved in water?
- What is meant by the term dissociation and what is an example of a substance that dissociates?
- What are the essential characteristics of a solution that is (include information regarding its behavior as well as the types of substances which indicate these types of solutions):
  - A nonconductor
  - A good conductor
  - A poor conductor

Procedure

**Part 1 - Conductivity Classification**

Test and record the conductivity of each substance and solution listed below using the method demonstrated by your instructor. Unless otherwise noted waste will go in the inorganic waste container. Then, after noting the range of conductivities measured, classify each as having essentially no ions, a few ions, or many ions. For each substance record also the major and minor species present in the sample. Your data table may be similar to the one shown below:

**TABLE 4.6**

Substance	Bond Type (Polar covalent, Ionic)	Conductivity	Ions (None, Few, Many)	Major species present in solution	Minor species in solution	Observations
HBr <sub>(aq)</sub>	polar covalent, fully ionized	Good	Many	H <sub>3</sub> O <sup>+</sup> <sub>(aq)</sub> , Br <sup>-</sup> <sub>(aq)</sub>	n/a	All three bulbs lit.
HF <sub>(aq)</sub>	polar covalent, partially ionized	Poor	Few	HF <sub>(aq)</sub>	H <sub>3</sub> O <sup>+</sup> <sub>(aq)</sub> , F <sup>-</sup> <sub>(aq)</sub>	Two bulbs lit weakly.
Deionized water						
Tap water						
Methanol, CH <sub>3</sub> OH <sub>(l)</sub> (organic waste)						
CH <sub>3</sub> OH <sub>(aq)</sub> (organic waste)						
Glacial acetic acid CH <sub>3</sub> CO <sub>2</sub> H <sub>(l)</sub>						
CH <sub>3</sub> CO <sub>2</sub> H <sub>(aq)</sub> (Add water slowly to glacial acetic acid and record how the conductivity changes.)						
Sucrose C <sub>12</sub> H <sub>22</sub> O <sub>11</sub> <sub>(s)</sub>						
C <sub>12</sub> H <sub>22</sub> O <sub>11</sub> <sub>(aq)</sub>						
NaCl <sub>(s)</sub>						
NaCl <sub>(aq)</sub>						
KClO <sub>3</sub> <sub>(s)</sub> (Test in a crucible.)						
KClO <sub>3</sub> <sub>(l)</sub> molten (Heat crucible using a Bunsen burner.)						
0.1 M HgCl <sub>2</sub>						
0.1 M HCl						
0.1 M NaOH						
0.1 M CH <sub>3</sub> CO <sub>2</sub> H						
0.1 M NH <sub>3</sub>						
0.1 M NaCl						

## Part 2 - Effect of Solvent

Test the conductivity of the following solvents and mixtures. Put all waste in the organic waste container. Your data table may be similar to the one shown below:

TABLE 4.7

Substance	Conductivity	Ions (None, Few, Many)	Major species present in solution	Minor species in solution	Observations
Xylene, C <sub>8</sub> H <sub>10</sub> (l) (premade)					
HCl in xylene (premade)					
Aqueous layer (after mixing water with HCl in xylene)					

## Part 3 - Correlating Chemical and Conductivity Behavior

Be sure to use similar amounts of each solid and acid. You will compare the conductivity of the acids with their reaction rates. Compare the rates of reaction (fast, medium, slow, or no reaction) of:

- calcium carbonate, CaCO<sub>3</sub>(s), with 6 M acetic acid and 6 M hydrochloric acid;
- zinc with 6 M acetic acid and 6 M hydrochloric acid.

TABLE 4.8

Acid	Conductivity Before rxn	Ions (None, Few, Many) Before rxn	Rxn Rate CaCO <sub>3</sub> (s)	Rxn Rate Zn(s)	Observations
HCl <sub>(aq)</sub>					
HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub> (aq)					

## Part 4 - Observing Changes in Conductivity for Ionic Reactions

Measure the conductivities of the acid and base alone, and then titrate the acid with base using a pasteur pipet and observe the changes in conductivity. Add several drops of phenolphthalein to the acid before you begin the titration. Continue adding several more drops of base after the phenolphthalein turns pink. Note how the conductivity changes as you add base to the acid. Note the changes in conductivity as well as any other changes that may occur as you perform the titration.

- Perform the experiment with 0.1 M HCl with 0.1 M NaOH;
- Perform the experiment with 0.1 M CH<sub>3</sub>COOH with 0.1 M NH<sub>3</sub>;
- Perform the experiment with 0.1 M H<sub>2</sub>SO<sub>4</sub> with 0.1 M Ba(OH)<sub>2</sub>.

TABLE 4.9

	Conductivity	Ions (None, Few, Many)	Observations
Before Rxn: reactant 1			
Before Rxn: reactant 2			



5. Explain the difference in behavior in the conductivity of  $\text{KClO}_3$  (s) and  $\text{KClO}_3$  (l).

**Part 2**

1. Why is the conductivity of HCl different in the two solvents?

2. What causes this difference?

**Part 3**

1. Write the equation for the reactions of:
- calcium carbonate and acetic acid
  
  
  
  
  
  
  
  
  
  
  - calcium carbonate and hydrochloric acid;
  
  
  
  
  
  
  
  
  
  
  - zinc and acetic acid;







*Post Lab Questions*

1. Listed below are several substances and their conductivities when dissolved in water. Based on this information, write the formulas for all of the individual species present. Omit water from your list.

**TABLE 4.10**

Substance	Conductivity	Major Species Present	Minor Species Present
HF <sub>(aq)</sub>	Poor	HF <sub>(aq)</sub>	H <sup>+</sup> <sub>(aq)</sub> , F <sup>-</sup> <sub>(aq)</sub>
K <sub>2</sub> SO <sub>4(aq)</sub>	Good		
CH <sub>2</sub> O <sub>(l)</sub>	Non-conductive		
HCHO <sub>2(aq)</sub>	Poor		

2. Recognition of ionic or molecular species present. For each substance, write the formula(s) of the principal species (molecular or ionic) present in major amounts in the aqueous solution if the substance is soluble; if it is only slightly soluble, use the molecular formula followed by (s). Also include the minor species where appropriate. The first three are done as examples.

**TABLE 4.11**

Substance	Major Species Present	Minor Species Present
BaSO <sub>4(s)</sub>	BaSO <sub>4(s)</sub>	-----
CH <sub>3</sub> COOH <sub>(aq)</sub>	CH <sub>3</sub> COOH <sub>(aq)</sub>	CH <sub>3</sub> COO <sup>-</sup> <sub>(aq)</sub> , H <sup>+</sup> <sub>(aq)</sub>
NaCl <sub>(aq)</sub>	Na <sup>+</sup> <sub>(aq)</sub> , Cl <sup>-</sup> <sub>(aq)</sub>	-----
Al(OH) <sub>3(s)</sub>		
Fe <sub>2</sub> (SO <sub>4</sub> ) <sub>3(aq)</sub>		
HBr <sub>(aq)</sub>		
H <sub>3</sub> PO <sub>4(aq)</sub>		
Mg(OH) <sub>2(s)</sub>		
NH <sub>4</sub> C <sub>2</sub> H <sub>3</sub> O <sub>2(aq)</sub>		
C <sub>2</sub> H <sub>5</sub> OH <sub>(aq)</sub>		
HgCl <sub>2(aq)</sub>		
CuS <sub>(s)</sub>		



c. Nitric acid and magnesium acetate

d. Ammonium chloride and sodium hydroxide

e. Barium chloride and calcium nitrate

f. Potassium hydrogen carbonate and sulfuric acid

g. Aluminum hydroxide and nitric acid

h. Aqueous ammonia and sulfuric acid

i. Magnesium nitrate and zinc chloride