# 7.7: Solubility Rules for Ionic Compounds

## Learning Objectives

- Define and give examples of electrolytes.
- Determine the solubility of common ionic compounds.

When some substances are dissolved in water, they undergo either a physical or a chemical change that yields ions in solution. These substances constitute an important class of compounds called **electrolytes**. Substances that do not yield ions when dissolved are called nonelectrolytes. If the physical or chemical process that generates the ions is essentially 100% efficient (all of the dissolved compound yields ions), then the substance is known as a strong electrolyte. If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, it is called a weak electrolyte.

Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (Figure 7.7.1).

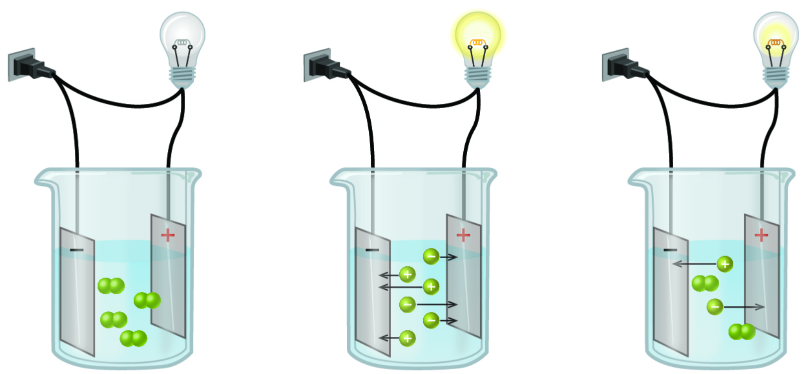


Figure 7.7.1: Solutions of nonelectrolytes, such as ethanol, do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte.

## Ionic Electrolytes

Water and other polar molecules are characterized by a slightly positive region and a slightly negative region and are therefore attracted to ions, as shown in Figure 7.7.2. The electrostatic attraction between an ion and a molecule with a dipole is called an ion-dipole attraction. These attractions play an important role in the dissolution of ionic compounds in water, which will be later discussed in Chapter 14.

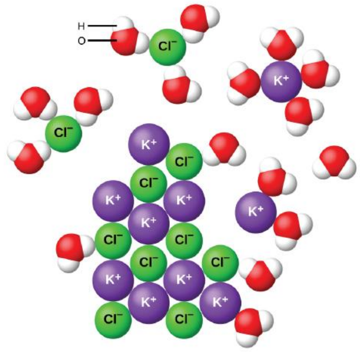


Figure 7.7.2: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the $K^+$ and $Cl^-$ ions. Water molecules in front of and behind the ions are not shown.

When ionic compounds dissolve in water, the ions in the solid separate and disperse uniformly throughout the solution because water molecules surround and solvate the ions, reducing the strong electrostatic forces between them. This process represents a physical change known as dissociation. Under most conditions, ionic compounds will dissociate nearly completely when dissolved, and so they are classified as strong electrolytes.

Let us consider what happens at the microscopic level when we add solid KCl to water. Ion-dipole forces attract the slightly positive (hydrogen) end of the polar water molecules to the negative chloride ions at the surface of the solid, and they attract the slightly negative (oxygen) end to the positive potassium ions. The water molecules penetrate between individual $K^+$ and $Cl^-$ ions and surround them, reducing the strong interionic forces that bind the ions together and letting them move off into solution as solvated ions, as Figure 7.7.2 shows. The reduction of the electrostatic attraction permits the independent motion of each hydrated ion in a dilute solution, resulting in an increase in the disorder of the system, as the ions change from their fixed and ordered positions in the crystal to mobile and much more disordered states in solution. This increased disorder is responsible for the dissolution of many ionic compounds, including KCl, which dissolve with absorption of heat.

In other cases, the electrostatic attractions between the ions in a crystal are so large, or the ion-dipole attractive forces between the ions and water molecules are so weak, that the increase in disorder cannot compensate for the energy required to separate the ions, and the crystal is insoluble. Such is the case for compounds such as calcium carbonate (limestone), calcium phosphate (the inorganic component of bone), and iron oxide (rust).

## Solubility Rules

Some combinations of aqueous reactants result in the formation of a solid precipitate as a product. However, some combinations will not produce such a product. If solutions of sodium nitrate and ammonium chloride are mixed, no reaction occurs. One could write an equation showing an exchange of ions; but both products, sodium chloride and ammonium nitrate, are soluble and remain in the solution as ions. It is useful to be able to predict when a precipitate will occur in a reaction. To do so, you can use a set of guidelines called **solubility rules** (Tables 7.7.1 and 7.7.2).

# Table 7.7.1: Solubility Rules for Soluble Substances

| Soluble in Water | Important Exceptions (Insoluble) |
|---|---|
| All Group IA (alkali metals) and ammonium compounds (containing $Li^+$, $Na^+$, $K^+$, $Rb^+$, $Cs^+$, $NH_4^+$) | none |
| All nitrates, chlorates, and acetates (containing $NO_3^-$, $ClO_3^-$, $C_2H_3O_2^-$) | none |

| Soluble in Water | Important Exceptions (Insoluble) |
|---|---|
| Most sulfates (containing $SO_4^{2-}$) | Sulfates of $Ca^{2+}$, $Sr^{2+}$, $Ba^{2+}$, $Ag^{+}$, and $Pb^{2+}$ |
| Most chlorides, bromides, and iodides (containing $Cl^-$, $Br^-$, $I^-$) | Chlorides, bromides, and iodides of $Ag^+$, $Pb^{2+}$, and $Hg_2^{2+}$ |

## Table 7.7.2: Solubility Rules for Insoluble Substances

| Insoluble in Water | Important Exceptions (Soluble) |
|---|---|
| Most carbonates, oxalates, and phosphates (containing $CO_3^{2-}$, $C_2O_4^{2-}$, $PO_4^{3-}$) | Group IA (alkali metals) and ammonium compounds |
| Most hydroxides (containing $OH^-$) | Group IA (alkali metals) and ammonium compounds<br><br>Hydroxides of $Ca^{2+}$, $Sr^{2+}$, $Ba^{2+}$ sparingly soluble – though usually considered insoluble |
| Most sulfides (containing $S^{2-}$) | Group IA (alkali metals) and ammonium compounds<br><br>Sulfides of $Ca^{2+}$, $Sr^{2+}$, $Ba^{2+}$ |

### ✅ Example 7.7.1: Solubility

Classify each compound as soluble or insoluble

A. $Zn(NO_3)_2$
B. $PbBr_2$
C. $Sr_3(PO_4)_2$

**Solution**

A. Nitrates are soluble in water with no exceptions, so $Zn(NO_3)_2$ is soluble.
B. Most bromides are soluble in water. However, combinations with $Pb^{2+}$ are an exception, so $PbBr_2$ is insoluble.
C. Most phosphates are insoluble and there is no exception when combined with $Sr^{2+}$, so $Sr_3(PO_4)_2$ is insoluble.

### ✏️ Exercise 7.7.1: Solubility

Classify each compound as soluble or insoluble.

A. $Mg(OH)_2$
B. KBr
C. $Pb(NO_3)_2$

**Answer A**

insoluble

**Answer B**

soluble

**Answer C**

soluble

## Summary

- Substances that dissolve in water to yield ions are called electrolytes.
- Nonelectrolytes are substances that do not produce ions when dissolved in water.
- Solubility rules allow prediction of what products will be insoluble in water.