

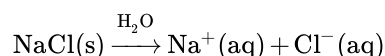
## 11.3: Ionic Equations - A Closer Look

### Learning Objectives

- Write ionic equations for chemical reactions between ionic compounds.
- Write net ionic equations for chemical reactions between ionic compounds.

For single-replacement and double-replacement reactions, many of the reactions included ionic compounds—compounds between metals and nonmetals, or compounds that contained recognizable polyatomic ions. Now, we take a closer look at reactions that include ionic compounds.

One important aspect about ionic compounds that differs from molecular compounds has to do with dissolution in a liquid, such as water. When molecular compounds, such as sugar, dissolve in water, the individual molecules drift apart from each other. When ionic compounds dissolve, *the ions physically separate from each other*. We can use a chemical equation to represent this process—for example, with NaCl:



When NaCl dissolves in water, the ions separate and go their own way in solution; the ions are now written with their respective charges, and the (aq) phase label emphasizes that they are dissolved (Figure 11.3.1). This process is called **dissociation**; we say that the ions *dissociate*.

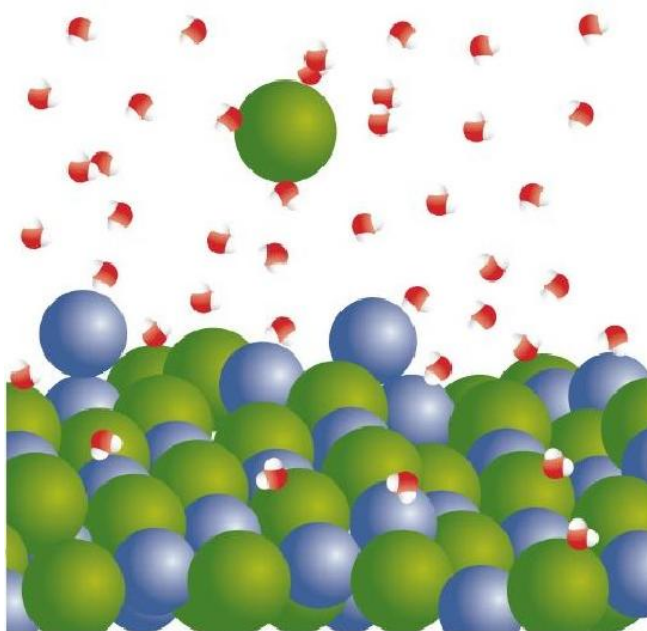
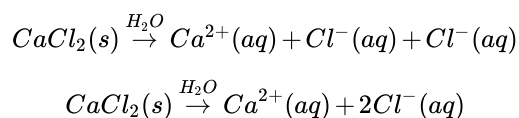


Figure 11.3.1 Ionic Solutions. When an ionic compound dissociates in water, water molecules surround each ion and separate it from the rest of the solid. Each ion goes its own way in solution.

All ionic compounds that dissolve behave this way. This behavior was first suggested by the Swedish chemist Svante August Arrhenius [1859–1927] as part of his PhD dissertation in 1884. Interestingly, his PhD examination team had a hard time believing that ionic compounds would behave like this, so they gave Arrhenius a barely passing grade. Later, this work was cited when Arrhenius was awarded the Nobel Prize in Chemistry. Keep in mind that when the ions separate, *all* the ions separate. Thus, when  $\text{CaCl}_2$  dissolves, the one  $\text{Ca}^{2+}$  ion and the two  $\text{Cl}^-$  ions separate from one another:



That is, the two chloride ions go off on their own. They do not remain as  $\text{Cl}_2$  (that would be elemental chlorine; these are chloride ions), and they do not stick together to make  $\text{Cl}_2^-$  or  $\text{Cl}_2^{2-}$ . They become dissociated ions in their own right. Polyatomic ions also retain their overall identity when they are dissolved.

### ✓ Example 11.3.1

Write the chemical equation that represents the dissociation of each ionic compound.

1.  $\text{KBr}$
2.  $\text{Na}_2\text{SO}_4$

#### Solution

1.  $\text{KBr(s)} \rightarrow \text{K}^+(\text{aq}) + \text{Br}^-(\text{aq})$
2. Not only do the two sodium ions go their own way, but the sulfate ion stays together as the sulfate ion. The dissolving equation is  $\text{Na}_2\text{SO}_4(\text{s}) \rightarrow 2\text{Na}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$

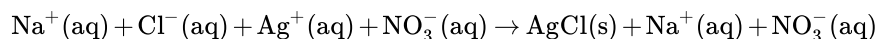
### ? Exercise 11.3.1

Write the chemical equation that represents the dissociation of  $(\text{NH}_4)_2\text{S}$ .

#### Answer



When chemicals in solution react, the proper way of writing the chemical formulas of the dissolved ionic compounds is in terms of the dissociated ions, not the complete ionic formula. A **complete ionic equation** is a chemical equation in which the dissolved ionic compounds are written as separated ions. Solubility rules are very useful in determining which ionic compounds are dissolved and which are not. For example, when  $\text{NaCl(aq)}$  reacts with  $\text{AgNO}_3(\text{aq})$  in a double-replacement reaction to precipitate  $\text{AgCl(s)}$  and form  $\text{NaNO}_3(\text{aq})$ , the complete ionic equation includes  $\text{NaCl}$ ,  $\text{AgNO}_3$ , and  $\text{NaNO}_3$  written as separate ions:



This is more representative of what is occurring in the solution.

### ✓ Example 11.3.1

Write the complete ionic equation for each chemical reaction.

1.  $\text{KBr(aq)} + \text{AgC}_2\text{H}_3\text{O}_2(\text{aq}) \rightarrow \text{KC}_2\text{H}_3\text{O}_2(\text{aq}) + \text{AgBr(s)}$
2.  $\text{MgSO}_4(\text{aq}) + \text{Ba(NO}_3)_2(\text{aq}) \rightarrow \text{Mg(NO}_3)_2(\text{aq}) + \text{BaSO}_4(\text{s})$

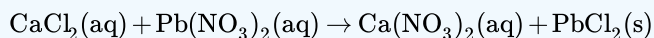
#### Solution

For any ionic compound that is aqueous, we will write the compound as separated ions.

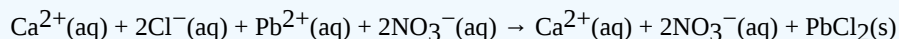
1. The complete ionic equation is  $\text{K}^+(\text{aq}) + \text{Br}^-(\text{aq}) + \text{Ag}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) \rightarrow \text{K}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) + \text{AgBr(s)}$
2. The complete ionic equation is  $\text{Mg}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{BaSO}_4(\text{s})$

### ? Exercise 11.3.1

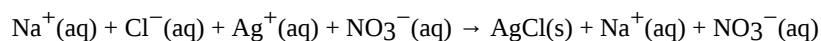
Write the complete ionic equation for



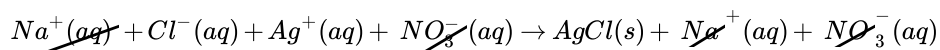
#### Answer



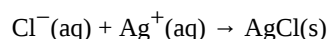
You may notice that in a complete ionic equation, some ions do not change their chemical form; they stay exactly the same on the reactant and product sides of the equation. For example, in



the  $\text{Ag}^+(\text{aq})$  and  $\text{Cl}^-(\text{aq})$  ions become  $\text{AgCl}(\text{s})$ , but the  $\text{Na}^+(\text{aq})$  ions and the  $\text{NO}_3^-(\text{aq})$  ions stay as  $\text{Na}^+(\text{aq})$  ions and  $\text{NO}_3^-(\text{aq})$  ions. These two ions are examples of spectator ions—ions that do nothing in the overall course of a chemical reaction. They are present, but they do not participate in the overall chemistry. It is common to cancel spectator ions (something also done with algebraic quantities) on the opposite sides of a chemical equation:



What remains when the spectator ions are removed is called the **net ionic equation**, which represents the actual chemical change occurring between the ionic compounds:



It is important to reiterate that the spectator ions are still present in solution, but they do not experience any net chemical change, so they are not written in a net ionic equation.

### ✓ Example 11.3.1

Write the net ionic equation for each chemical reaction.

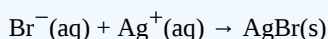
- $\text{K}^+(\text{aq}) + \text{Br}^-(\text{aq}) + \text{Ag}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) \rightarrow \text{K}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) + \text{AgBr}(\text{s})$
- $\text{Mg}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{BaSO}_4(\text{s})$

#### Solution

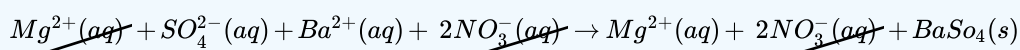
- In the first equation, the  $\text{K}^+(\text{aq})$  and  $\text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$  ions are spectator ions, so they are canceled:



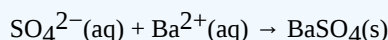
The net ionic equation is



- In the second equation, the  $\text{Mg}^{2+}(\text{aq})$  and  $\text{NO}_3^-(\text{aq})$  ions are spectator ions, so they are canceled:

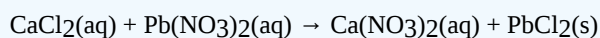


The net ionic equation is

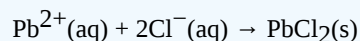


### ? Exercise 11.3.1

Write the net ionic equation for



#### Answer



### ✓ Chemistry is Everywhere: Soluble and Insoluble Ionic Compounds

The concept of solubility versus insolubility in ionic compounds is a matter of degree. Some ionic compounds are very soluble, some are only moderately soluble, and some are soluble so little that they are considered insoluble. For most ionic compounds, there is also a limit to the amount of compound that can be dissolved in a sample of water. For example, you can dissolve a maximum of 36.0 g of NaCl in 100 g of water at room temperature, but you can dissolve only 0.00019 g of AgCl in 100 g of water. We consider NaCl soluble but AgCl insoluble.

One place where solubility is important is in the tank-type water heater found in many homes in the United States. Domestic water frequently contains small amounts of dissolved ionic compounds, including calcium carbonate ( $\text{CaCO}_3$ ). However,

$\text{CaCO}_3$  has the relatively unusual property of being less soluble in hot water than in cold water. So as the water heater operates by heating water,  $\text{CaCO}_3$  can precipitate if there is enough of it in the water. This precipitate, called *limescale*, can also contain magnesium compounds, hydrogen carbonate compounds, and phosphate compounds. The problem is that too much limescale can impede the function of a water heater, requiring more energy to heat water to a specific temperature or even blocking water pipes into or out of the water heater, causing dysfunction.



Figure 11.3.2 Tank water heater © Thinkstock. Most homes in the United States have a tank-type water heater like this one.

Another place where solubility versus insolubility is an issue is the Grand Canyon. We usually think of rock as insoluble. But it is actually ever so slightly soluble. This means that over a period of about two billion years, the Colorado River carved rock from the surface by slowly dissolving it, eventually generating a spectacular series of gorges and canyons. And all because of solubility!



Figure 11.3.3: The Grand Canyon was formed by water running through rock for billions of years, very slowly dissolving it. Note the Colorado River is still present in the lower part of the photo. (Sonaal Bangera via [unsplash](https://unsplash.com/photos/Grand-Canyon))

## Key Takeaways

- Ionic compounds that dissolve separate into individual ions.
- Complete ionic equations show dissolved ionic solids as separated ions.
- Net ionic equations show only the ions and other substances that change in a chemical reaction.

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