An aqueous solution is one that is occurring in water. What makes water significant is that it can allow for substances to dissolve and/or be dissociated into ions within it.

**Electrolytes**

Water is generally the solvent found in aqueous solution, where a solvent is the substance that dissolves the solute. The solute is the substance or compound being dissolved in the solvent. A solute has fewer number of particles than a solvent, where it’s particles are in random motion. Interestingly, aqueous solutions with ions conduct electricity to some degree. Pure water, having a very low concentration of ions, cannot conduct electricity. When a solute dissociates in water to form ions, it is called an electrolyte, due to the solution being a good electrical conductor. When no ions are produced, or the ion content is low, the solute is a non-electrolyte. Non-electrolytes do not conduct electricity or conduct it to a very small degree. In an aqueous solution a strong electrolyte is considered to be completely ionized, or dissociated, in water, meaning it is soluble. Strong acids and bases are usually strong electrolytes. A weak electrolyte then is considered to be one that is not completely dissociated, therefore still containing whole compounds and ions in the solution. Weak acids and bases are generally weak electrolytes. In other words, strong electrolytes have a better tendency to supply ions to the aqueous solution than weak electrolytes, and therefore strong electrolytes create an aqueous solution that is a better conductor of electricity.

- Most soluble ionic compounds and few molecular compounds are strong electrolytes.
- Most molecular compounds are weak or non electrolytes.

Example \((\PageIndex{1})\)

Here’s an example of MgCl\(_2\) in water:

\[\text{MgCl}_2 \rightarrow \text{Mg}^{2+}_{(aq)} + 2\text{Cl}^-_{(aq)}\]

The ionic compound dissociates completely to form ions in water, therefore, it is a strong electrolyte.

Now let’s look at a weak electrolyte:

\[\text{HC}_2\text{H}_3\text{O}_2(aq) \rightleftharpoons \text{H}^+_{(aq)} + \text{C}_2\text{H}_3\text{O}_2^-_{(aq)}\]

The ionic compound, \((\text{HC}_2\text{H}_3\text{O}_2)\) in this situation, only partially dissociates, as expressed by the double arrows in the reaction. This means that the reaction is reversible and never goes to completion.

The \((\text{H}^+)\) cation is a proton that interacts with the \((\text{H}_2\text{O})\) molecules that it is submerged in. The interaction is called hydration. The actual \(\text{H}^+\) ion does not exist in the aqueous solution. It is the hydronium ion, \((\text{H}_3\text{O}^+)\) that interacts with water to create additional species like \((\text{H}_5\text{O}_2^+)\), \((\text{H}_9\text{O}_4^+)\), and \((\text{H}_7\text{O}_3^+)\). \((\text{H}_3\text{O}^+)\) can simply be described as the hydration of one \(\text{H}^+\) and one water molecule. For nonelectrolytes, all that needs to be done is write the molecular formula because no reaction or dissociation occurs. One example of a nonelectrolyte is sugar: written as \((\text{C}_6\text{H}_{12}\text{O}_6)\).
Ion Concentrations

In an aqueous solution the amount of ions of a species is related to the number of moles of that species per concentration of the substance in the aqueous solution. **Molarity** is the number of moles of a solute \((\text{n})\) divided by the total volume \((\text{V})\) of the solution:

\[
[M] = \frac{\text{n}}{\text{V}}
\]

Molarity, or concentration, can also be represented by placing the solute within brackets (e.g., \([\text{Cl}^-]\)) for the concentration of chloride ions).

Example \((\PageIndex{2})\)

Determine the concentration of \(\text{K}^+\) in an aqueous solution of 0.238 M \(\text{KNO}_3\).

**SOLUTION**

Since there is one mole of potassium in \(\text{KNO}_3\), multiply the concentration of the species by the number of moles of the atom to obtain:

\[
[[\text{K}^+] = (0.238\; \text{M} \; \text{KNO}_3) \times (1 \; \text{mol} \; \text{K}^+)= 0.238\; \text{M}
\]

Although not asked, there is also one mole of nitrite ions in one mole of \(\text{KNO}_3\), so its concentration is also 0.238 M:

\[
[[\text{NO}_3^-] = (0.238\; \text{M} \; \text{KNO}_3) \times (1 \; \text{mol} \; \text{NO}_3^-)= 0.238\; \text{M}
\]

**FOLLOWUP**

The stoichiometry always dictates the concentration, which was a simple 1:1 ratio for \(\text{KNO}_3\). However, for more complex situations, different ratios will be encountered. For instance, if consider the dissolving of \(\text{Al}_2(\text{SO}_4)_3\):

\[
\text{Al}_2(\text{SO}_4)_3 \rightarrow 2 \text{Al}^{3+}_{(aq)} + 3 \text{SO}^{2-}_{4(aq)}
\]

If the concentration of \(\text{Al}_2(\text{SO}_4)_3\) is 0.019 M, what is the concentration of \(\text{Al}^{3+}_{(aq)}\)? Simply multiply 0.019 M by the stoichiometric factor of \(\text{Al}^{3+}_{(aq)}\) in \(\text{Al}_2(\text{SO}_4)_3\), which is 2:1. The concentration of \(\text{Al}^{3+}_{(aq)}\) then becomes 0.038 M:

\[
[[\text{Al}^{3+}] = (0.019 \; \text{M} \; \text{Al}_2(\text{SO}_4)_3) \times (2 \; \text{mol} \; \text{Al}^{3+})= 0.038\; \text{M}
\]

Although not asked, the concentration of \(\text{SO}_4^{2-}\) is 0.057 M via the same argument;

\[
[[\text{SO}_4^{2-}]= (0.019 \; \text{M} \times 3 \; \text{mol} \; \text{SO}_4^{2-}) = 0.057\; \text{M}
\]
Precipitation Reactions

Precipitation reactions occur when the resulting product of an aqueous solution is insoluble. This means that there is a solid produced, called the **precipitate**. The precipitate is a combination of cation and anions forming an ionic bond. Precipitates are used in manufacturing chemicals, so that certain ions can be isolated by forming precipitates with them. We can predict if a precipitate is produced using the **solubility rules** for common ionic solids:

1. Salts of **Group 1** cations (with some exceptions for \((\text{Li}^+))\) and the \((\text{NH}_4^+)\) cation are soluble.
2. Nitrates, acetates, and perchlorates are soluble.
3. Salts of silver, lead, and mercury(I) are insoluble.
4. Chlorides, bromides, and iodides are soluble.
5. Carbonates, phosphates, sulfides, oxides, and hydroxides are insoluble (sulfides of **Group 2** cations and hydroxides of \((\text{Ca}^{2+})\), \((\text{Sr}^{2+})\), and \((\text{Ba}^{2+})\) are slightly soluble).
6. Sulfates are soluble except for those of calcium, strontium, and barium.

Example \((\text{PageIndex}[3])\)

Here's an example of a precipitation reaction:

\[
\text{AgNO}_3(aq) + \text{NaI}(aq) \rightarrow \text{AgI}(s) + \text{NaNO}_3(aq)
\]

This is considered the "whole" reaction, because all of the species involved are recognized. However, in an aqueous solution, the particles look more like this:

\[
\text{Ag}^+(aq) + \text{NO}_3^-(aq) + \text{Na}^+(aq) + \text{I}^-(aq) \rightarrow \text{AgI}(s) + \text{Na}^+(aq) + \text{NO}_3^-(aq)
\]

This is called the "ionic" form, because all the ions in the solution are shown. If the ions that are not involved in creating the solid are removed (the **spectator ions**), a net ionic equation is generated:

\[
\text{Ag}^+(aq) + \text{I}^-(aq) \rightarrow \text{AgI}(s)
\]

What are the spectator ions in this reaction?

\[
\text{NaOH}(aq) + \text{MgCl}_2(aq) \rightarrow \text{NaCl}(aq) + \text{Mg(OH)}_2(s)
\]

As you can see, it is **critical** to include the symbols for each species to identify what state they're in: either gaseous (g), solid (s), liquid (l) or aqueous (aq). It is also important to remember that these types of ionic equations also must have a balanced charge. You can use stoichiometric coefficients to ensure that both sides of the equation have equal net charges.

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Acid Base Reactions

An **acid** is a substance that gives off \((\text{H}^+)\) ions in an aqueous solution, while a **base** gives of \(\text{OH}^-\) ions. Strong acids almost completely dissociate to become \((\text{H}^+)\) ions, and strong bases dissociate to become \(\text{OH}^-\). There are only a few
strong acids and bases. Most are weak, meaning they produce few \((H^+\text{)}\) or \(OH^-\) ions in aqueous solution. Here is a list of common strong acids and bases:

<table>
<thead>
<tr>
<th>Strong Acids</th>
<th>Strong Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>LiOH</td>
</tr>
<tr>
<td>HBr</td>
<td>NaOH</td>
</tr>
<tr>
<td>HI</td>
<td>KOH</td>
</tr>
<tr>
<td>HClO₄</td>
<td>RbOH</td>
</tr>
<tr>
<td>HNO₃</td>
<td>CsOH</td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>Ca(OH)₂</td>
</tr>
<tr>
<td></td>
<td>Sr(OH)₂</td>
</tr>
<tr>
<td></td>
<td>Ba(OH)₂</td>
</tr>
</tbody>
</table>

In a **neutralization reaction** an acid and a base are combined to produce water and an aqueous salt. The acid and base neutralize/balance each other to get a result that is the salt. For example, determine which is the acid, base, and salt in this neutralization reaction?

\[
\text{HCl}_{(aq)} + \text{NaOH}_{(aq)} \rightarrow \text{NaCl}_{(aq)} + \text{H}_2\text{O}_{(l)}
\]

**References**


**Problems**

Predict if a reaction is likely to occur, and the resulting product:

a. \(\text{HI}_{(aq)} + \text{Zn(NO}_3{)_2}_{(aq)} \rightarrow\)

b. \(\text{CuSO}_4_{(aq)} + \text{Na}_2\text{CO}_3_{(aq)} \rightarrow\)

c. \(\text{AgNO}_3_{(aq)} + \text{CuCl}_2_{(aq)} \rightarrow\)

Determine the concentration:

- \([\text{Al}^{3+}]\) in 0.078 M \(\text{Al}_2(\text{SO}_4)_3\)

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