1.1: Solubility

Solution

The **solution** is a homogeneous mixture of two or more substances.

> **Solution related terminologies**
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> - **Miscible** substances make a solution upon mixing in any proportion. For example, ethanol and water are miscible to each other.
> - **Immiscible** substances do not make solutions upon mixing in any proportion.
> - **Partially miscible** substances can make a solution upon mixing up to a certain extent but not in all proportions.
> - **A solvent** is a larger substance in the solution.
> - **A solute** is a small amount of a substance in the solution.
> - **An unsaturated solution** is one in which the solvent holds solute less than the maximum limit, i.e., more solute can be dissolved.
> - **A saturated solution** is one in which the solvent holds the maximum amount of solute it can dissolve.

Water -a universal solvent

Water is one of the most important solvents because it is present all around us -it covers more than 70% of the earth and more than 60% of our body mass. Water is a polar molecule having a partially negative end on oxygen and a partially positive end on hydrogen atoms. that can dissolve most of the polar and ionic compounds. In ionic compounds, cations are held by anions through electrostatic interaction. When an ionic compound dissolves into water, it dissociates into cations and anions, each surrounded by a layer of water molecules held by ion-dipole interactions. The water molecules around ions make ion-dipole interaction by orienting their partial negative end towards cations and their partial positive end towards anions. The energy needed to break ion-ion interaction in the ionic compounds is partially compensated by the energy released by establishing the ion-dipole interactions. The energy gained due to ion-dipole interactions, and nature's tendency to disperse is the driving forces responsible for the dissolution of ionic compounds.

> **Solubility**
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> **Solubility** is the ability of a substance to form a solution with another substance.

The solubility of a solute in a specific solvent is quantitatively expressed as the concentration of the solute in the saturated solution. Usually, the solubility is tabulated in the units of grams of solute per 100 mL of solvent (g/100 mL). The solubility of ionic compounds in water varies over a wide range. All ionic compounds dissolve to some extent.

> For practical purposes, a substance is considered **insoluble** when its solubility is less than 0.1 g per 100 mL of solvent.

For example, lead(II)iodide ($\mathrm{PbI_2}$) and silver chloride (AgCl) are insoluble in water because the solubility of $\mathrm{PbI_2}$ is 0.0016 mol/L of the solution and the solubility of AgCl is about 1.3 x 10^{-5} mol/L of solution. Potassium iodide (KI) and $\mathrm{Pb(NO_3)_2}$ are soluble in water. When aqueous solutions of KI and $\mathrm{Pb(NO_3)_2}$ are mixed, the insoluble combination of ions, i.e., $\mathrm{PbI_2}$ in this case, precipitates, as illustrated in Figure 1.1.1.

Figure 1.1.1: Precipitation reaction: $Pb(NO_3)_2(aq) + 2\,KI(aq) \longrightarrow PbI_2(s)\downarrow + 2\,KNO_3(aq)$. source: PRHaney [CC BY-SA (https://creativecommons).

Solubility guidelines for dissolution of ionic compounds in water

There are no fail-proof guidelines for predicting the solubility of ionic compounds in water. However, the following guidelines can predict the solubility of most ionic compounds.

Soluble ions

1. Salts of alkali metals (Li^+, Na^+, K^+, Rb^+, Cs^+) and ammonia (NH_4^+) are soluble. For example, $NaCl$, and $(NH_4)_3PO_3$ are soluble.
2. Salts of nitrate (NO_3^-), acetate (CH_3COO^-), and perchlorate (ClO_4^-) are soluble. For example, $Pb(NO_3)_2$, and $Ca(CH_3COO)_2$ are soluble.
3. Salts of chloride (Cl^-), bromide (Br^-), and Iodide (I^-) are soluble, except when the cation is Lead (Pb^{2+}), Mercury (Hg_2^{2+}), or Silver (Ag^+). Remember the acronym "LMS" based on the first letter of the element name or phrase 'Let Me See" to recall Lead, Mercury, and Silver.
4. Sulfates (SO_4^{2-}) are soluble except when the cation is, Pb^{2+}, Hg_2^{2+}, or Ag^+ (recall "Let Me See" for Lead, Mercury, and Silver) or a heavy alkaline earth metal ion: calcium (Ca^{2+}), barium (Ba^{2+}), or strontium (Sr^{2+}). (Remember the acronym "CBS" based on the first letter of the element name, or phrase "Come By Soon" to recall calcium, barium, and strontium.)

Insoluble ions

1. Hydroxide (OH^-) and sulfides (S^{2-}) are insoluble except when the cation is a heavy alkaline earth metal ion: Ca^{2+}, Ba^{2+}, or Sr^{2+} (recall "Come By Soon" for calcium, barium, and strontium), alkali metals and ammonium. For example, $Mg(OH)_2$ and CuS are insoluble.
2. Carbonates (CO_3^{2-}), phosphates (PO_4^{3-}), and oxide (O^{2-}) are insoluble except when the cation is an alkali metal ion or ammonium. For example, $CaCO_3$, and Fe_2O_3 are insoluble.
3. If there is a conflict between the two guidelines, the first listed guideline has priority. For example, $CaCO_3$ is insoluble (rule#6), but Na_2CO_3 is soluble (rule#1 has priority over rule#6).

Precipitation reactions

Precipitation reactions are a class of chemical reactions in which two solutions are mixed, and a solid product, called a precipitate, separates out. Precipitation reaction happening upon mixing solutions of ionic compounds in water can be predicted as illustrated in Figure 1.1.2. The first step is to list the soluble ionic compounds and then cross-combine the cations of one with the anion of the other to make the potential products. If any of the potential products are insoluble ionic compounds, they will precipitate out. For example, when $NaOH$ solution is mixed with $MgCl_2$ solution, $Mg(OH)_2$ is a cross-combination that forms an insoluble compound, it will precipitate out.

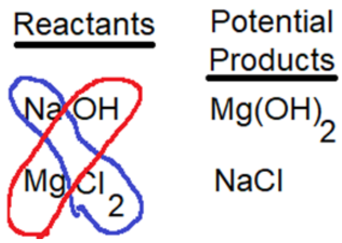

Figure 1.1.2: Cross-combine the cation-anion in the reactants. If any of the cross-combination is an insoluble salt, it will precipitate out, e.g.: $NaOH(aq) + MgCl_2(aq) \longrightarrow Mg(OH)_2(s)\downarrow + NaCl(aq)$.

Figure 1.1.3 shows precipitates of some insoluble ionic compounds formed by mixing aqueous solutions of appropriate soluble ionic compounds.

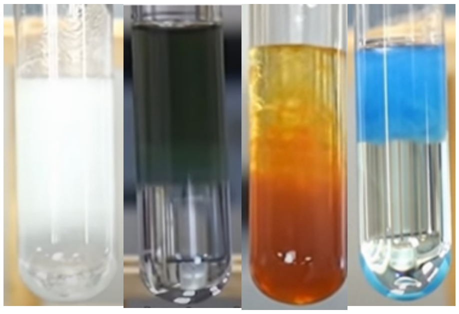

Figure 1.1.3: The precipitates of some insoluble ionic compounds formed by mixing the aqueous solution of appropriate soluble ionic compounds. The precipitates are from left: white Calcium sulfate ($CaSO_4$), black Iron(II) hydroxide ($Fe(OH)_2$), brown Iron(III) hydroxide ($Fe(OH)_3$), and blue Copper(II) hydroxide ($Cu(OH)_2$). Note that the precipitate is not yet settled at the bottom of the solution; it is still in suspension form in these examples. Source: https://youtu.be/jltLlzZ6FqU
