

5.2: Solubility

Solubility and its related terminologies

The majority of solutes do not dissolve in water or other solvents in all proportions.

What is solubility

The maximum proportion of the solute that can dissolve in a given amount of the solvent, usually expressed in grams of solute in 100 grams of solvent, is called the **solubility** of the solute in the solvent.

1. Substances that make a solution when mixed in any proportion are called **miscible**. For example, ethanol and water are miscible.
2. Some substances make a solution when mixed in some proportion but not in all proportions; these are called **partially miscible**. For example, n-butanol can mix in water up to 7.3 g n-butanol/100 ml water.
3. Substances that almost do not dissolve in each other are called **immiscible**. For example, n-Hexane is immiscible in water.
4. The solution that has not yet reached its solubility limits and can dissolve more solute added is called an **unsaturated solution**.
5. The solution that has reached its solubility limits and can not dissolve if more solute is added to it is called a **saturated solution**.

When a solute is added to a saturated solution, it does dissolve, but, at the same time, the dissociated components recombine to form the crystals of the solute, i.e., recrystallize, at the same rate, so that there is no net dissolution, as illustrated in Fig. 5.2.1.

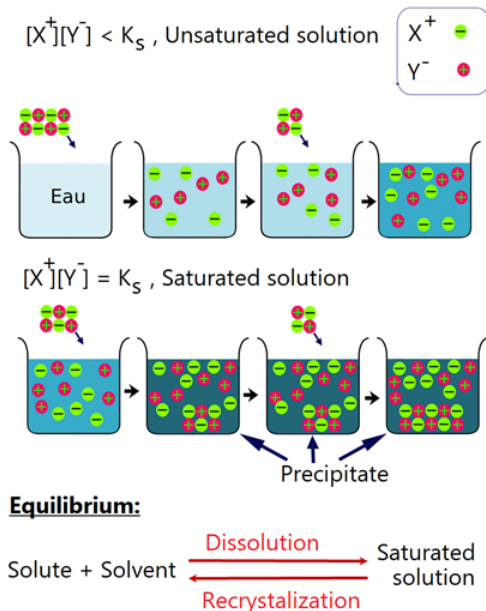


Figure 5.2.1: Illustration of unsaturated and saturated solution and the dynamic equilibrium between dissolution and recrystallization in the saturated solution. Source: No machine-readable author provided. Romary assumed (based on copyright claims). / CC BY-SA (<http://creativecommons.org/licenses/by-sa/3.0/>)

Effect of temperature on solubility

The solubility of most of the solids and liquids increases as the temperature increases.

For example, when a saturated solution of sugar in water is heated, it can dissolve more sugar. Fig 5.2.2 show the solubility vs temperature curves for some compounds.

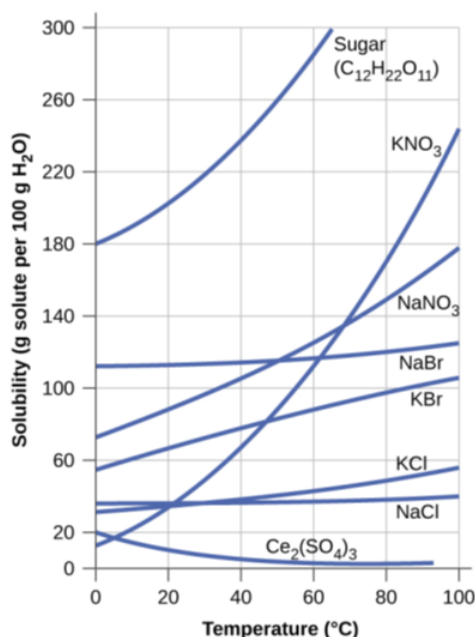


Figure 5.2.2: Solubility of some ionic and polar molecular compounds in water as a function of temperature. Download for free at <https://openstax.org/details/books/chemistry>.

When a hot saturated solution is cooled, the concentration of solute in the solution becomes above the solubility limits, making a **supersaturated solution**.

The supersaturated solution is unstable and ultimately crystallizes out the excess solute leaving any impurities in the solution, as illustrated in Fig. 5.2.3. This process is called re-crystallization, which is used to purify the solutes.

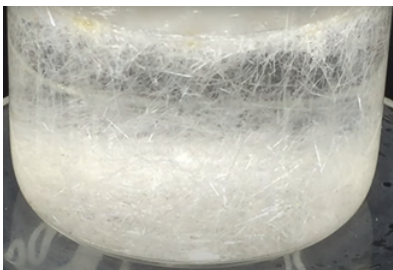


Figure 5.2.3: Salicylic acid recrystallized in water. Source: <https://www.youtube.com/watch?v=QwqvzfPneUo>

Medical issues related with solubility

The crystallization of excess solute from a supersaturated solution is responsible for some medical problems like gout and kidney stone. Gout is the crystallization of uric acid in the cartilage, tendons, and soft tissues when the concentration of the uric acid in blood plasms exceeds its solubility limit of ~7 mg/100 mL at 37 °C. It causes redness, swelling, and pain in the affected area, as illustrated in Fig. 5.2.4. Kidney stones are solid materials formed in the urinary tubes, as illustrated in Fig. 5.2.5. Kidney stones are the result of the crystallization of excess calcium phosphate, calcium oxalate, or uric acid in the urine.



Figure 5.2.4: Illustration of gout as a result of the recrystallization of uric acid in the cartilage, tendons, and soft, causes swelling, joint stiffness, and aching around the joint in the foot. Source: www.scientificanimations.com/ / CC BY-SA (<https://creativecommons.org/licenses/by-sa/4.0>)

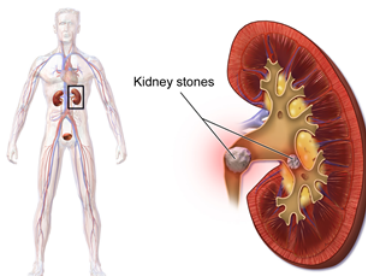


Figure 5.2.5: Kidney Stones illustrated. Source: Blausen.com staff (2014). "Medical gallery of Blausen Medical 2014". WikiJournal of Medicine 1 (2). DOI:10.15347/wjm/2014.010. ISSN 2002-4436/ CC BY (<https://creativecommons.org/licenses/by/3.0>)

Opposite to the solids and liquids, the solubility of gases generally decreases with an increase in temperature.

That is why the carbonated water releases dissolved gas when heated, causing pressure increase, which, in turn, causes the bursting of the soda can

Effect of pressure on solubility

The pressure has almost no effect on the solubility of solids and liquids but has a strong effect on the solubility of gases.

Henry's law

The solubility of gases in liquids is directly proportional to the pressure of the gas above the liquid.

An increase in pressure causes a decrease in the gas volume that increases the gas concentration. More frequent collision of the gas molecules with the gas-liquid boundary in a concentrated solution causes an increase in the rate of dissolution of the gas in the liquid, as illustrated in Fig 5.2.6. The opposite happens when the gas pressure decreases. For example, carbon dioxide starts bubbling out when a soda can is open because the gas escapes resulting in a decrease in the gas pressure above the liquid and a decrease in the solubility of the gas in water.

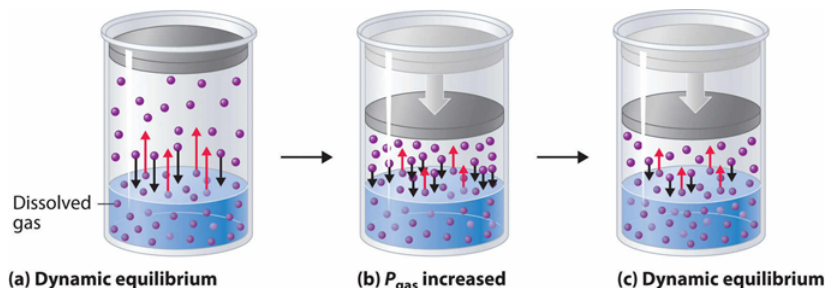


Figure 5.2.6: The Solubility of a gas increases as the partial pressure increases at a constant temperature. Source: Abozenadah, H., Bishop, A., Bittner, S. and Flatt, P.M. (2017) Preparatory Chemistry. CC BY-NC-SA. Available at: <https://wou.edu/chemistry/courses/on...ory-chemistry/>

Solubility guidelines for dissolution of ionic compounds in water

If the solubility of a compound is less than 0.01 mol/L, it is considered **insoluble**.

The solubility of ionic compounds in water depends on the nature of the compound. For example, lead(II)iodide (PbI_2) and silver chloride (AgCl) are insoluble in water because the solubility of PbI_2 is 0.0016 mol/L of the solution and the solubility of AgCl is about 1.3×10^{-5} mol/L of solution. Potassium iodide (KI) and $\text{Pb}(\text{NO}_3)_2$ are soluble in water. When aqueous solutions of KI and $\text{Pb}(\text{NO}_3)_2$ are mixed, the concentration of PbI_2 in the mixture goes above its solubility limits, and it precipitates out, as illustrated in Fig. 5.2.7.



Figure 5.2.7: Precipitation reaction: $\text{Pb}(\text{NO}_3)_2(\text{aq}) + 2\text{KI}(\text{aq}) \rightarrow \text{PbI}_2(\text{s})\downarrow + 2\text{KNO}_3(\text{aq})$. source: PRHaney [CC BY-SA (<https://creativecommons.org/licenses/by-sa/3.0/>)]

There are no fail-proof guidelines for predicting the solubility of ionic compounds in water. However, the following guideline 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. No exceptions.
2. Salts of nitrate (NO_3^-), acetate (CH_3COO^-), and perchlorate (ClO_4^-) are soluble. No exceptions.
3. Salts of chloride (Cl^-), bromide (Br^-), and Iodide (I^-) are soluble, except when the cation is Pb^{2+} , Hg_2^{2-} , or 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 Pb^{2+} , Hg_2^{2-} , Ag^+ , or a heavy alkaline earth metal ions: calcium (Ca^{2+}), barium (Ba^{2+}), or strontium (Sr^{2+}). ((recall “Let Me See” for Lead, Mercury, and Silver. Remember the acronym “CBS” based on the first letter of the element name, or the 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 an alkali metal, ammonia, or a heavy alkaline earth metal ions: Ca^{2+} , Ba^{2+} , and Sr^{2+} . (Recall the phrase “Come By Soon” to recall calcium, barium, and strontium.)
2. Carbonates (CO_3^{2-}), phosphates (PO_4^{3-}), and oxide (O^{2-}) are insoluble except when the cation is an alkali metal, ammonia
3. If there is a conflict between the two guidelines, then the guideline listed first has priority. For example, the salts of insoluble ions become soluble when the cation is an alkali metal, ammonia (rule#1).

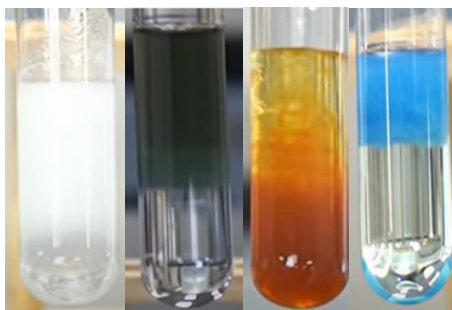


Figure 5.2.8: The precipitates of some insoluble ionic compounds formed by mixing the aqueous solution of appropriate soluble ionic compounds. The precipitates are from the left: white Calcium sulfate (CaSO_4), black Iron(II) hydroxide ($\text{Fe}(\text{OH})_2$), brown Iron(III) hydroxide ($\text{Fe}(\text{OH})_3$), and blue Copper(II) hydroxide ($\text{Cu}(\text{OH})_2$). Source: <https://youtu.be/jltLzZ6FqU>

Fig. 5.2.8 shows precipitates of some insoluble ionic compounds formed by mixing aqueous solutions of appropriate soluble ionic compounds. The precipitation can be predicted, as illustrated in Fig. 5.2.9. List the ions of 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 is an insoluble ionic compound, it precipitates out.

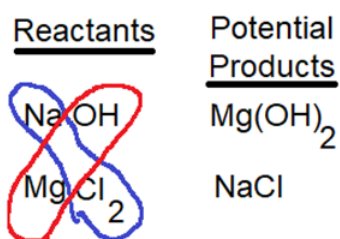


Figure 5.2.9: Cross-combine the cation-anion in the reactants. If any of the cross-combination is an insoluble salt, it will precipitate out, e.g: $\text{NaOH}(\text{aq}) + \text{MgCl}_2(\text{aq}) \rightarrow \text{Mg}(\text{OH})_2(\text{s}) \downarrow + \text{NaCl}(\text{aq})$

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