

Teachers guide: Solubility Rules

There are a number of patterns in the data obtained from measuring the solubility of different salts. These patterns form the basis for the rules outlined in the table below, which can guide predictions of whether a given salt will dissolve in water. These rules are based on the following definitions of the terms *soluble*, *insoluble*, and *slightly soluble*.

- A salt is soluble if it dissolves in water to give a solution with a concentration of at least 0.1 moles per liter at room temperature.
- A salt is insoluble if the concentration of an aqueous solution is less than 0.001 M at room temperature.
- Slightly soluble salts give solutions that fall between these extremes.

Solubility Rules for Ionic Compounds in Water

Soluble Salts

1. The Na^+ , K^+ , and NH_4^+ ions form *soluble salts*. Thus, NaCl , KNO_3 , $(\text{NH}_4)_2\text{SO}_4$, Na_2S , and $(\text{NH}_4)_2\text{CO}_3$ are soluble.
2. The nitrate (NO_3^-) ion forms *soluble salts*. Thus, $\text{Cu}(\text{NO}_3)_2$ and $\text{Fe}(\text{NO}_3)_3$ are soluble.
3. The chloride (Cl^-), bromide (Br^-), and iodide (I^-) ions generally form *soluble salts*. Exceptions to this rule include salts of the Pb^{2+} , Hg_2^{2+} , Ag^+ , and Cu^+ ions. ZnCl_2 is soluble, but CuBr is not.
4. The sulfate (SO_4^{2-}) ion generally forms *soluble salts*. Exceptions include BaSO_4 , SrSO_4 , and PbSO_4 , which are insoluble, and Ag_2SO_4 , CaSO_4 , and Hg_2SO_4 , which are slightly soluble.

Insoluble Salts

1. Sulfides (S^{2-}) are usually *insoluble*. Exceptions include Na_2S , K_2S , $(\text{NH}_4)_2\text{S}$, MgS , CaS , SrS , and BaS .
2. Oxides (O^{2-}) are usually *insoluble*. Exceptions include Na_2O , K_2O , SrO , and BaO , which are soluble, and CaO , which is slightly soluble.
3. Hydroxides (OH^-) are usually *insoluble*. Exceptions include NaOH , KOH , $\text{Sr}(\text{OH})_2$, and $\text{Ba}(\text{OH})_2$, which are soluble, and $\text{Ca}(\text{OH})_2$, which

is slightly soluble.

4. Chromates (CrO_4^{2-}) are usually *insoluble*. Exceptions include Na_2CrO_4 , K_2CrO_4 , $(\text{NH}_4)_2\text{CrO}_4$, and MgCrO_4 .

5. Phosphates (PO_4^{3-}) and carbonates (CO_3^{2-}) are usually *insoluble*. Exceptions include salts of the Na^+ , K^+ , and NH_4^+ ions.

Discussion:

The heating (or cooling) effect that is observed when some substances are dissolved is called the heat of solution. This is another physical property that characterizes a compound. The heat of solution is the difference between the amount of energy associated with the formation of a layer of solvent molecules around the solute and the amount of energy associated with formation of the crystal packing in the solute. Another effect that you might have noted is the speed of dissolving. This property is dependent on particle size, that is, the surface area exposed to solvent compared to the mass of the particle. The calcium chloride we are using is a fine powder (with some larger clumps) while salt is composed of larger crystalline particles. As you saw if you did the optional exercise, solubility is affected by temperature. In general, substances are more soluble at higher temperatures. Some compounds have unusual solubility profiles. Sodium sulfate, for instance, reaches a maximum solubility in water at 33°C and becomes less soluble at higher temperatures. Salt also is slightly less soluble in hot water than at room temperature. Thus, our method of determining solubility is a bit crude, since we are not controlling for temperature or taking account of the dissolving rate.

To determine the limit of solubility we added a substance to a solvent until no more would dissolve. At this point we had a saturated solution. The concentration of the solute in the solvent was at the maximum possible for a stable mixture of the two. Here stable refers to a condition in which the mixture does not change over time, as long as temperature remains constant and physical changes like evaporation of the solvent or chemical changes like decomposition of the solute do not occur. The point of maximum concentration is called saturation. Concentration can be measured in g of solute/mL of solvent (in effect the density of the solute in the solvent).

Supplemental Experiments:

The optional experiment in the previous section showed that the density of a substance was changed by adding salt. The optional experiment

above showed that you can change solubility by changing the temperature. The following experiment shows that you can recover a dissolved solid from a liquid by means other than evaporation - by changing its solubility.

Create a nearly-saturated solution of salt in 5 mL of water (add slightly less than the amount you used to determine the solubility above). Add 5 mL of the alcohol, liquid A, to the solution and stir well. What happens? Recall that salt is not very soluble in alcohol. By changing the nature of the solvent by adding another liquid, the solubility of the salt was reduced. The change that occurred was a precipitation, but not one caused by a reaction, rather one caused by a change in the property of the solvent. You may want to filter the precipitate, let it dry, and test its solubility to convince yourself that the salt was recovered unchanged (except the particles may be smaller size).

In this section, we have concentrated on the solubility of solids in liquids. In the next section, we will see that one can make a solution of any phase of matter in any of the same or other phases. Try to determine the solubility of alcohol in water. Start with 3.0 g in 5.0 mL of water. Add 1.0 g of alcohol. If it is soluble, it will disappear into the water and you will not see two liquid layers as you did with the effect of salt on density experiment above. Add another 1.0 g of alcohol to the solution. Does it all dissolve? Add another 1.0 g of alcohol. Does this last portion dissolve? Now there is more alcohol than water. Which substance is the solvent and which is the solute? If you kept at this experiment, you would soon find that any amount of alcohol will dissolve in any amount of water and vice versa. Alcohol and water are thus said to be **miscible**. Any two substances that exhibit this behavior can be described as miscible, although the term is normally applied to liquids.

Effect of Temperature on Solubility:

The solubility of solutes is dependent on temperature. When a solid dissolves in a liquid, a change in the physical state of the solid analogous to melting takes place. Heat is required to break the bonds holding the molecules in the solid together. At the same time, heat is given off during the formation of new solute -- solvent bonds.

CASE I: Decrease in solubility with temperature:

If the heat given off in the dissolving process is greater than the heat required to break apart the solid, the net dissolving reaction is exothermic (energy given off). The addition of more heat (increases temperature) inhibits the dissolving reaction since excess heat is already being produced by the reaction. This situation is not very common where an increase in temperature produces a decrease in solubility.

CASE II: Increase in solubility with temperature:

If the heat given off in the dissolving reaction is less than the heat required to break apart the solid, the net dissolving reaction is endothermic (energy required). The addition of more heat facilitates the dissolving reaction by providing energy to break bonds in the solid. This is the most common situation where an increase in temperature produces an increase in solubility for solids.

The use of first-aid instant cold packs is an application of this solubility principle. A salt such as ammonium nitrate is dissolved in water after a sharp blow breaks the containers for each. The dissolving reaction is endothermic - requires heat. Therefore the heat is drawn from the surroundings, the pack feels cold.

Solubility of Gases vs. Temperature:

The variation of solubility for a gas with temperature can be determined by examining the graphic on the left.

As the temperature increases, the solubility of a gas decrease as shown by the downward trend in the graph .

More gas is present in a solution with a lower temperature compared to a solution with a higher temperature.

The reason for this gas solubility relationship with temperature is very similar to the reason that vapor pressure increases with temperature. Increased temperature causes an increase in kinetic energy. The higher kinetic energy causes more motion in molecules which break intermolecular bonds and escape from solution.

This gas solubility relationship can be remembered if you think about what happens to a "soda pop" as it stands around for awhile at room temperature. The taste is very "flat" since more of the "tangy" carbon dioxide bubbles have escaped. Boiled water also tastes "flat" because all of the oxygen gas has been removed by heating.

Quiz: Thermal pollution is merely waste heat that has been transferred to water or air. How is the concentration of dissolved oxygen in water be effected by thermal pollution?

Gas Pressure and Solubility:

Liquids and solids exhibit practically no change of solubility with changes in pressure. Gases as might be expected, increase in solubility with an increase in pressure. Henry's Law states that: The solubility of a gas in a liquid is directly proportional to the pressure of that gas above the surface of the solution.

If the pressure is increased, the gas molecules are "forced" into the solution since this will best relieve the pressure that has been applied. The number of gas molecules is decreased. The number of gas molecules dissolved in solution has increased as shown in the graphic on the left.

Carbonated beverages provide the best example of this phenomena. All carbonated beverages are bottled under pressure to increase the carbon dioxide dissolved in solution. When the bottle is opened, the pressure above the solution decreases. As a result, the solution effervesces and some of the carbon dioxide bubbles off.

Quiz: Champagne continues to ferment in the bottle. The fermentation produces CO₂. Why is the cork wired on a bottle of champagne?

Deep sea divers may experience a condition called the "bends" if they do not readjust slowly to the lower pressure at the surface. As a result of

breathing compressed air and being subjected to high pressures caused by water depth, the amount of nitrogen dissolved in blood and other tissues increases. If the diver returns to the surface too rapidly, the nitrogen forms bubbles in the blood as it becomes less soluble due to a decrease in pressure. The nitrogen bubbles can cause great pain and possibly death.

To alleviate this problem somewhat, artificial breathing mixtures of oxygen and helium are used. Helium is only one-fifth as soluble in blood as nitrogen. As a result, there is less dissolved gas to form bubbles.