

Today we will practice the skill of THINKING LIKE A CHEMIST while considering the concept of solubility. Platinum stars will be on the line.

Consider the following demonstration, and describe macroscopically.

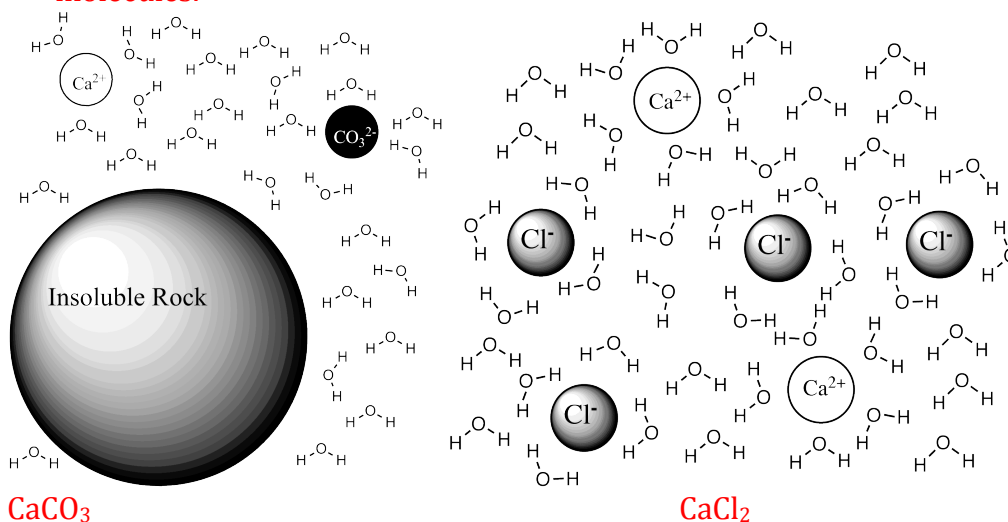
1. Limestone (CaCO_3) is placed in a beaker of water. Salt (CaCl_2) is placed in a beaker of water.

CaCO_3 remains solid and is unchanged. CaCl_2 dissolved and “disappears.”

2. Describe the demonstration from a microscopic perspective using words and a picture.

CaCO_3 perhaps dissolves some very slight amount, however the bulk remains insoluble.

CaCl_2 dissolves completely into its ions which are surrounded by water molecules.



3. Use chemical equations to model the changes.



Be prepared to explain if called upon in class.

The extent of solubility is given by the “solubility” of soluble solutes. For example, the solubility of CaCl_2 at room temperature is listed on Wikipedia as 64.7 g/ 100 g water, whereas the solubility of NaCl is listed as 35.72 g/ 100 g water.

4. Which salt is more soluble CaCl_2 or NaCl on a mass percent basis?

CaCl_2 is more soluble since more can dissolve in that same amount of water

5. Express both of those solubilities in terms of molar solubilities (express in units of moles per liter of solution). At room temperature, the density of a saturated solution of CaCl_2 is 1.435 g/ml and the density of a saturated solution of NaCl is 1.199 g/ml. Which salt is more soluble in terms of molar solubility?

Molar Mass CaCl_2 : 111.07 g/mol

Molar Mass NaCl : 58.5 g/mol

g_{solute} = mass of solute
 g_{solvent} = mass of solvent

d_{solution} = density of solution
 mm_{solute} = molar mass of solute

Molar Solubility CaCl_2 :

$$x = \left(\frac{g_{\text{solute}} d_{\text{solution}} * 1000}{(g_{\text{solute}} + g_{\text{solvent}}) * mm_{\text{solute}}} \right) = \left(\frac{(64.7 \text{ g}) * (1.435 \text{ g/ml}) (1000 \text{ ml/L})}{(64.7 \text{ g} + 100 \text{ g}) * 111.07 \text{ g/mol}} \right) \approx 5.1 \text{ M}$$

Molar Solubility NaCl :

$$x = \left(\frac{g_{\text{solute}} d_{\text{solution}} * 1000}{(g_{\text{solute}} + g_{\text{solvent}}) * mm_{\text{solute}}} \right) = \left(\frac{(35.72 \text{ g}) * (1.199 \text{ g/ml}) (1000 \text{ ml/L})}{(35.72 \text{ g} + 100 \text{ g}) * 58.5 \text{ g/mol}} \right) \approx 5.4 \text{ M}$$

NaCl is more soluble than CaCl_2 because $5.4 \text{ M} > 5.1 \text{ M}$

For insoluble compounds, the extent of solubility is given by something called the ion product, K_{sp} . The ion product is the product of the molar concentrations of all the ions in the formula unit. The ion product for CaCO_3 is $K_{\text{sp}} = [\text{Ca}^{2+}][\text{CO}_3^{2-}]$. You can write ion products for any salt (soluble or insoluble). The ion product for CaCl_2 is

$$K_{\text{sp}} = [\text{Ca}^{2+}][\text{Cl}^{-1}][\text{Cl}^{-1}] \quad \text{or} \quad K_{\text{sp}} = [\text{Ca}^{2+}][\text{Cl}^{-1}]^2$$

You can determine the molar solubility for a salt if the K_{sp} value is given. Or, if you know the molar concentrations of at least one ion in the solution, you can calculate the K_{sp} of the salt! The ion product is an equilibrium condition. That is, the K_{sp} is a mathematical expression of the concentrations of the ions in solution at the point of saturation.

6. The K_{sp} value for CaCO_3 is 8.7×10^{-9} . Calculate the concentration of the calcium ion at equilibrium using the relationship $K_{sp} = [\text{Ca}^{2+}][\text{CO}_3^{2-}]$.

$$8.7 \times 10^{-9} = [\text{Ca}^{2+}][\text{CO}_3^{2-}]$$

Since Ca^{2+} and CO_3^{2-} are in a 1:1 ratio (from the chemical formula, CaCO_3)

We know: $[\text{Ca}^{2+}] = [\text{CO}_3^{2-}]$, so

$$8.7 \times 10^{-9} = [\text{Ca}^{2+}][\text{Ca}^{2+}] \text{ or } 8.7 \times 10^{-9} = [\text{Ca}^{2+}]^2$$

$$\sqrt{8.7 \times 10^{-9}} = \sqrt{[\text{Ca}^{2+}]^2}$$

$$9.3 \times 10^{-5} = [\text{Ca}^{2+}]$$

7. From the equilibrium concentration of Ca^{2+} ion, determine the molar solubility for CaCO_3 remembering the stoichiometric ratio that the calcium ion is in a one to one ratio with the formula unit.

Since $[\text{Ca}^{2+}] = 9.3 \times 10^{-5}$ and $[\text{Ca}^{2+}] = \text{moles of CaCO}_3$ that dissolve according to the equation:



Molar solubility for CaCO_3 is 9.3×10^{-5}

8. Just to practice another example – try to calculate the molar solubility of AgCl . The $K_{sp} = 1.8 \times 10^{-10}$.

$$1.8 \times 10^{-10} = [\text{Ag}^+][\text{Cl}^-]$$

$$1.8 \times 10^{-10} = x^2$$

$$\sqrt{1.8 \times 10^{-10}} = \sqrt{x^2}$$

$$x = 1.34 \times 10^{-5}$$

Dr. VDB or LaB will now give you a mini lecture on how you can determine which compound is more or less soluble based on the value of K_{sp} .

To understand the power of the concept of the K_{sp} , you must be able to understand aqueous chemistry from a chemist's point of view.

9. Watch the demonstration of the precipitation reaction and describe it from a macroscopic point of view.

Two clear solutions are poured together and the solution turns cloudy with a precipitate

10. Given that the two solutions were ammonium carbonate and calcium chloride, write the chemical equation that describes the reaction.

