Department of Chemistry
University of Texas at Austin

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Physical Equilibria Unit Activity – Solubility KEY

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

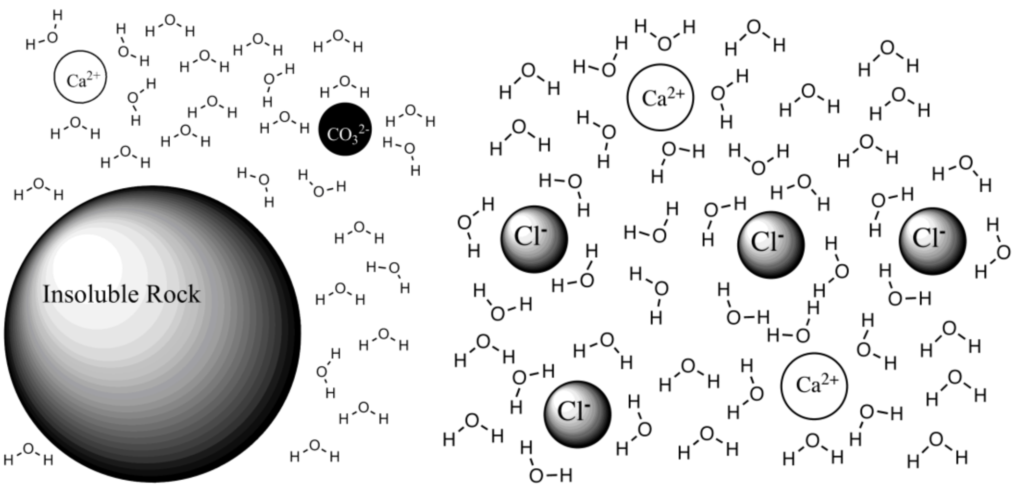
Consider the following demonstration, and describe macroscopically.

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.



$CaCO_3$ $CaCl_2$

3. Use chemical equations to model the changes.

$CaCO_3 \rightarrow Ca^{2+}(aq) + CO_3^{2-}(aq)$

$CaCl_2 \rightarrow Ca^{2+}(aq) + 2Cl^{-}(aq)$

Be prepared to explain if called upon in class.

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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 solubilites (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 $CaCl_2$:

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

Molar Solubility NaCl:

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

NaCl is more soluble than $CaCl_2$ because 5.4M > 5.1M

For insoluble compounds, the extent of solubility is given by something called the ion product, Ksp. 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 Ksp = $[Ca^{2+}][CO_3^{2-}]$. You can write ion products for any salt (soluble or insoluble). The ion product for $CaCl_2$ is

Ksp = $[Ca^{2+}][Cl^{-1}][Cl^{-1}]$ or Ksp = $[Ca^{2+}][Cl^{-1}]^2$

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

6. The Ksp value for $CaCO_3$ is 8.7 X 10^{-9}. Calculate the concentration of the calcium ion at equilibrium using the relationship Ksp = $[Ca^{2+}][CO_3^{2-}]$.

8.7 X 10^{-9} = $[Ca^{2+}][CO_3^{2-}]$
Since Ca^{2+} and CO_3^{2-} are in a 1:1 ratio (from the chemical formula, $CaCO_3$)
We know: $[Ca^{2+}] = [CO_3^{2-}]$, so
8.7 X 10^{-9} = $[Ca^{2+}][Ca^{2+}]$ or 8.7 X 10^{-9} = $[Ca^{2+}]^2$
$\sqrt{8.7X10^{-9}} = \sqrt{[Ca^{2+}]^2}$
9.3 X 10^{-5} = $[Ca^{2+}]$

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 $[Ca^{2+}]$ = 9.3 X 10^{-5} and $[Ca^{2+}] = [CaCO_3]$ according to the equation:
$CaCO_3 \rightarrow Ca^{2+}(aq) + CO_3^{2-}(aq)$

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

8. Just to practice another example – try to calculate the molar solubility of AgCl. The Ksp = 1.8 x 10^{-10}.
   1.8 x 10^{-10} = $[Ag^+][Cl^-]$
   1.8 x 10^{-10} = x^2
   $\sqrt{1.8x10^{-10}} = \sqrt{x^2}$
   x = 1.34 x 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 Ksp.

To understand the power of the concept of the Ksp, 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

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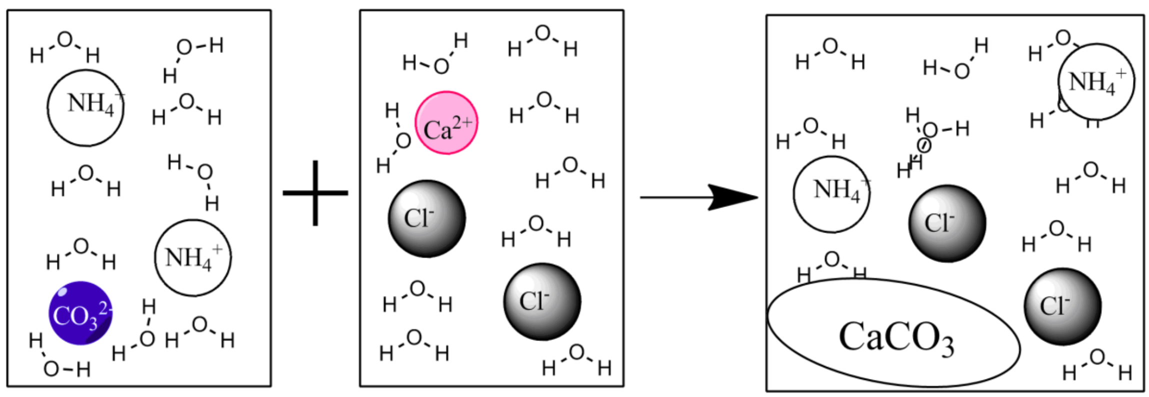
10. Given that the two solutions were ammonium carbonate and calcium chloride, write the chemical equation that describes the reaction.

$(NH_4)_2CO_3$ (aq) + $CaCl_2$ (aq) → $CaCO_3$ (s) + $2NH_4Cl$ (aq)

11. Draw microscopic view of the reactants and the products.



12. Write out the total ionic equation, which is the shorthand way to model the pictures you have drawn in #9.

$2NH_4^+$ (aq) + CO_3^{2-} (aq) + Ca^{2+} (aq)+ $2Cl^-$ (aq) → $CaCO_3$ (s) + $2NH_4^+$ (aq) + $2Cl^-$ (aq)

13. Write out the net ionic equation, which is how one demonstrates the species that are changing across the reaction. Label in your diagram from # 9 those species that did not change with the word "spectator ion".

$2NH_4^+$ (aq) + CO_3^{2-} (aq) + Ca^{2+} (aq) + $2Cl^-$ (aq) → $CaCO_3$ (s) + $2NH_4^+$ (aq) + $2Cl^-$ (aq)
^ ^ ^ ^
spectator spectator spectator spectator

Be prepared to share your answers with the class if called upon.

14. Practice the skill of writing the formula unit, total ionic and net ionic equations for the reaction of mixing a solution of lead II nitrate with a solution of potassium chromate.

Formula Unit: $Pb(NO_3)_2$ (aq) + K_2CrO_4 (aq) → $PbCrO_4$ (s) + $2KNO_3$ (aq)

Total: Pb^{2+}(aq) + $2NO_3^-$(aq) + $2K^+$(aq)+CrO_4^{2-}(aq) → $PbCrO_4$(s)+$2NO_3^-$(aq) + $2K^+$(aq)

Net: Pb^{2+} (aq) + CrO_4^{2-} (aq) → $PbCrO_4$ (s)