CHE 1302

Basic Principles of Modern Chemistry II

Week 11

Hello and welcome to the weekly resources for Chemistry 1302! This resource covers topics typically taught by professors during the 11^{th} week of classes.

On our website, https://baylor.edu/tutoring, you'll find the following links:

"Online Study Guide Resources" – If you don't see the topics you're learning right now, click here to find the weekly resources for the rest of the semester!

"How to Participate in Group Tutoring" - See if there is a Chemistry 1302 group tutoring session being hosted this semester – these are weekly question/answer sessions taught by our master tutors!

You can also view tutoring times for your course or schedule a private 30-minute appointment! Check out the website to learn more. You can also give us a call at (254)710-4135, or drop in. Our hours are Monday-Thursday 9 am – 8 pm on class days.

KEY WORDS: Solubility Equilibria, Solubility Product, pH Effects on Solubility, Precipitation

TOPIC OF THE WEEK: Solubility Equilibria & The Solubility Product

In Chem I, we learned about combinations of ions that are both soluble and insoluble. It turns out that the concept of insolubility isn't totally accurate. Everything is at least a *little* bit soluble, but sometimes the concentration of ions that can be tolerated in solution is very low, so only a few ions can leave the solid at a time. Most of it remains solid, which leads us to call it a **precipitate**.

What tool do we have that measures the concentrations on each side of an equilibrium? That's right—K expressions! The equilibrium constant expression for a **slightly soluble solid** in equilibrium with its ions is called the solubility product constant (K_{sp}). For the reaction $CaF_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2F^{-}_{(aq)}$, $K_{sp} = [Ca^{2+}][F^{-}]^2$

Notice that K_{sp} only depends on the ions in solution; it is a measure of the **concentrations of these particular ions** that the **solution can hold** *at a **given temperature***.

Solubility, on the other hand, is the maximum amount of solute dissolved in a solvent at equilibrium (commonly expressed in mol/L or g/L). If we were to make an I/C/E table, for the $CaF_{2(s)}$ dissociation, the solubility would be the initial concentration of $CaF_{2(s).}$ It's the amount of solid that we can put into a solution before it stops dissolving.

Note: The solubility product constant (K_{sp}) is a constant value (at a given temperature), while solubility is an equilibrium position which is subject to variation (e.g. in the presence of a common ion).

Say that you were **given Ksp** (these values are determined experimentally and listed in tables) and asked to **find solubility**. How are the two related mathematically? Here's a helpful video (~8 min): https://www.youtube.com/watch?v=WjiXbemBXkE

It's an I/C/E table problem. The initial concentration of each ion is 0 M. The solid is not used in calculation of Ksp, so it need not be considered. How much does each ion change in concentration? We're not sure, so put "x" as a placeholder. Be sure to adjust for stoichiometry as appropriate! (Review Week 6). Now, write an equation for Ksp. Remember,

$$K_c = \frac{[C]^m * [D]^n}{[A]^j * [B]^k}$$

The denominator, since the only reactant is a solid, is 1. If you've been given Ksp, the only unknown is x. Solve for x; this is the concentration of one of the ions. One mole of ions is stoichiometrically equivalent to one mole of solid, so **x=the solubility.**

Highlight 1: pH & Solubility

Outside factors can change the initial concentrations of the ions; when this happens, solubility changes. Review Le Chatelier's principle (Week 5): If a reactant or product is added or subtracted from an equilibrium, the equilibrium will shift to try to make up for the gain/loss.

Decrease in ion concentration:

Ions can act as acids/bases (review Week 9). If the pH (hydronium concentration) is changed, the equilibriums in which the ions act as acids/bases will shift. Ex: if the anion is a conjugate base of a weak acid: $HA + H_2O \rightleftharpoons A^- + H_3O^+$, increasing hydronium concentration will shift the equilibrium to the left, using up the anion as it goes. This increases solubility, because the solution can now tlerate more of the ions. Examples of such anions include: OH^- , S^{2-} , CO_3^{2-}, and CrO_4^{2-}. If the anion is a conjugate base of a strong acid, increasing hydronium concentration will have no effect, because the hydronium is not strong enough, compared to HA, to protonate the base.

Increase in ion concentration:

When an ion is added, the equilibrium will shift away from the ion, towards the solid. This common ion effect results in a decrease in solubility. Addition of OH^-, for example, increases pH and decreases solubility. Keep in mind: Ksp stays the same, but solubility can change.

Highlight 2: Precipitation and Qualitative Analysis

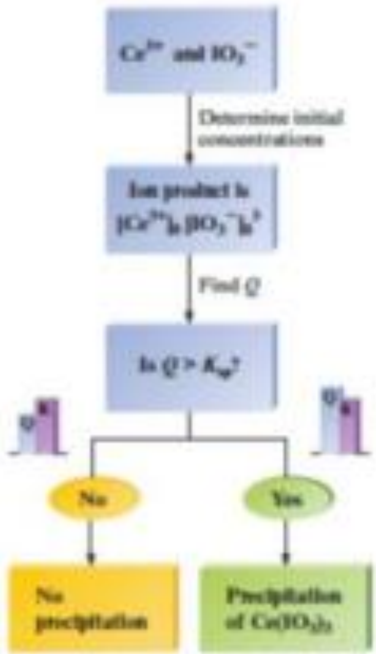
The reaction quotient, Q, and the solubility product constant, K_{sp}, can be compared to determine if precipitation of a given salt will occur in solution. We are comparing the **current proportion of ions in solution (Q)** to the **maximum acceptable proportion of ions** in solution (K_{sp}).

Q > Ksp: too many ions in solution → precipitation occurs

Q ≤ Ksp: no precipitation

Don't forget: The reaction quotient, Q, is the equilibrium constant expression using initial concentration (not equilibrium concentration).

As an example, the figure below shows how to determine if precipitation of Ce(IO3)3 will occur in a solution containing Ce3 + and IO3 - ions:



Picture from the textbook, *Chemistry, An Atoms First Approach by Zumdahl and Zumdahl.*

Check Your Understanding

1. 1 L of saturated $Ag_2CO_{3(aq)}$ contains 0.035 g of solute. What is the value of K_{sp}? What mass of Ag_2CO_3 is contained in 3.65 L of saturated solution?
2. Calculate the molar solubility of $PbBr_2$ in 0.1 M $Pb(NO_3)_2$.
3. When 200 mL of 1.7×10^{-4} M $Ba(NO_3)_2$ and 150 mL of 1.2×10^{-4} M Na_2SO_4 are mixed, will a precipitate form? K_{sp} for $BaSO_4$ is 1.1×10^{-10}

Things You May Struggle With

1. Not knowing the difference between solubility and solubility product constant.
2. Comparing Ksp values for salts with different total number of ions to determine the relative solubilities of the salts.
3. Forgetting to include the concentration of common ion present as initial concentration when trying to get the solubility of a salt present in a solution with a common ion.
4. In calculating the equilibrium concentration after precipitation occurs, failure to only use the excess ion present after precipitation as the initial concentration in the ICE method.
5. Not including the effect of dilution when mixing solution to get precipitates.
6. To get the formation constant expression for a complex ion formation, students sometimes forget to put the complex as a product and leave it as a reactant.

That's all this week! Please reach out if you have any questions and don't forget to visit the Tutoring Center website for further information at www.baylor.edu/tutoring. Answers to Check Your Understanding are below.

1. $K_{sp} = 8.8 \times 10^{-12}$; 0.13g Ag_2CO_3
2. 3.42×10^{-3}
3. Yes