CHE 1302

Basic Principles of Modern Chemistry II

Week 12

Hello and welcome to the weekly resources for Chemistry 1302! This resource covers topics typically taught by professors during the 12th 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.

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KEY WORDS: Solubility Product Constant Applied, Complex Ions

TOPIC OF THE WEEK: Solubility Product Constant Applied

General problem-solving strategy (see Week 11 for more specific Ksp I/C/E table tips)

1. Write balanced reaction equation
2. Create I/C/E table
3. Write K_{sp} equation
4. Plug in known values (you will be given either K_{sp} or solubility. Remember, solubility is the number of moles of solid that can be dissolved and is equal to x).
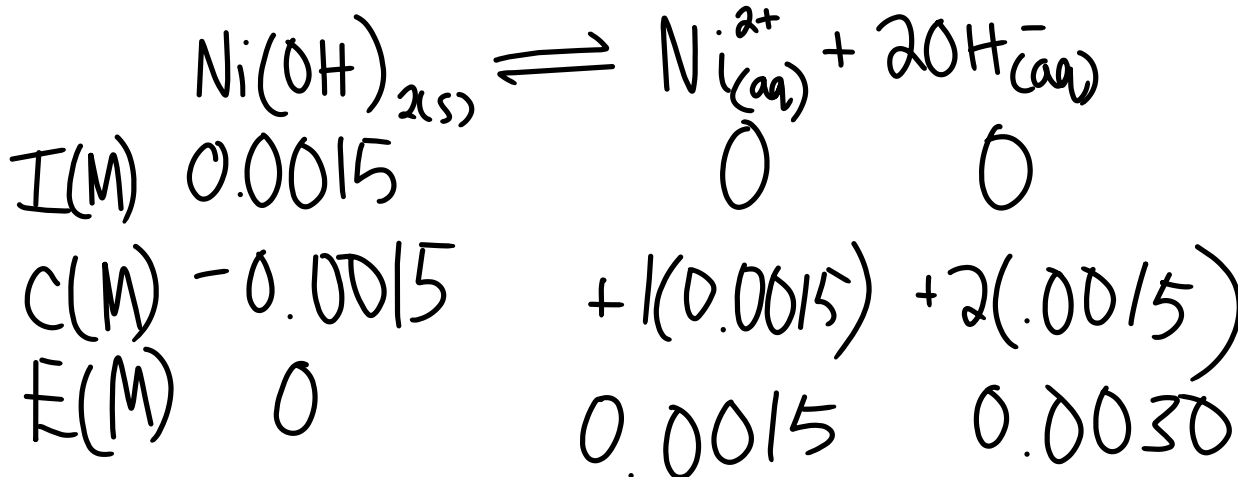5. Solve for unknowns.

Here's a worked example:

Approximately 0.14 g nickel(II) hydroxide, $Ni(OH)_{2(s)}$, dissolves per liter of water at 20°C. Calculate K_{sp} for $Ni(OH)_{2(s)}$ at this temperature.

1. Which part is the cation, and which is the anion? See Week 11 for tips on figuring this out. Nickel is a Group II metal, which means it'll want to lose 2 electrons in order to have a full valence shell. Losing two negatively charged electrons makes it Ni^{2+}. Hydroxide is a

polyatomic that should be memorized; it has a -1 charge, and there are two of them. The equation is $Ni(OH)_{2(s)} \rightleftharpoons Ni^{2+}_{(aq)} + 2OH^-_{(aq)}$

2. Now, create an I/C/E table. The ==solubility== was given (in g/L) in the problem. This means that, if 0.14 g of nickel hydroxide is added initially, that much will dissolve, making its final concentration zero. But we need to fill in the I/C/E table in terms of molarity, because that way we can relate it to the amount of product formed.
   a. Divide the mass by its molar mass (92.708g/mol) → 0.14 g/L of nickel hydroxide equals 0.0015 mol/L
   b. Initial concentrations of the ions will be 0 M.

$$Ni(OH)_{2(s)} \rightleftharpoons Ni^{2+}_{(aq)} + 2OH^-_{(aq)}$$

	$Ni(OH)_{2(s)}$	$Ni^{2+}_{(aq)}$	$2OH^-_{(aq)}$
I(M)	0.0015	0	0
C(M)	−0.0015	+1(0.0015)	+2(.0015)
E(M)	0	0.0015	0.0030

3. Write the solubility product constant (K_{sp}) expression and substitute the equilibrium concentration obtained from step 3 to get the answer; $K_{sp} = [Ni^{2+}][OH^-]^2 = 1.35*10^{-8}$
4. Notice: the unique thing about this problem is that we start with the solubility, not initial concentrations. Make sure you put each piece of info in the correct row! Another example – where does pH go? First, it needs to be converted to hydronium concentration, and then it can go in the equilibrium row.

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Highlight 1: Complex Ions

A complex ion is made of multiple elements, one of which is a metal. We know that metals tend to form positively charged ions; since the metal is positively charged, it attracts "ligands," or atoms with available electrons. The number of "ligands" that it attracts depends on how many empty orbitals it has. *Unless your prof tells you to do so, don't worry about being able to count the empty orbitals, as the formulas will be given!*

Some, but not all, of the molecules in solution will form the complex ion. The ratio of complex ion to "un-complexed" material is known as K_f.

You will often see situations in which molecules in solution can either form a precipitate or form a complex ion. Here is a way to approach these problems:

Ex. K_{sp} for AgCl is 1.6 x 10^{-10}, and K_f for the complex ion Ag(NH$_3$)$^{2+}$ is 1.7 x 10^7. Calculate the molar solubility of AgCl in 1.0 M NH$_3$.

1) The first equilibrium constant describes the dissolution of AgCl, and the second equilibrium constant describes the complex ion formation that uses Ag. Write out the equations for each.
   a. $AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)} \quad K_{sp} = 1.6 * 10^{-10}$
   b. $Ag^+_{(aq)} + NH_{3(aq)} \rightleftharpoons Ag(NH_3)^+_{(aq)} \quad K_f = 1.7 * 10^7$

2) Combine the equations. If they are added together, the Ag$^+$ will cancel out.
   a. $AgCl_{(s)} + NH_{3(aq)} \rightleftharpoons Cl^-_{(aq)} + Ag(NH_3)^+_{(aq)}$

3) Now, set up an I/C/E table. Use 1.0 M as the initial concentration of ammonia.

I   1.0 M   0   0
C   -x   +x   +x
E   1.0-x   x   x

4) Write the equilibrium constant expression for the I/C/E table equation. The combined Ks are multiplied to get the K for the combined equation.
   a. $K = K_{sp} * K_f = \dfrac{[Ag(NH_3)^+_3][Cl^-]}{[NH_3]}$

5) Fill in equation and solve for unknowns:
   a. $K = \dfrac{x^2}{1.0-x} = (1.6 * 10^{-10}) * (1.7 * 10^7) \rightarrow x = 5.0 * 10^{-2} M$

6) The concentration determined is the same as that of AgCl, *because:*
   a. *The Ag+ dissolved in solution is the same silver as that of the AgCl, assuming that all of the AgCl dissolved! And, if we're calculating solubility, the amount added will be exactly enough to dissolve all the way.*
   b. *And the [Ag(NH3)+] is the same as the [Ag+], because there will be plenty of ammonia to react with all of the Ag+ (Ag+ is limiting).*
   c. *So, the answer is $5.0 * 10^{-2} M$!*

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Highlight 2: Precipitation

Another type of problem that uses solubility constants will ask you to determine whether or not a precipitate is formed. In this case, you will need to (1) write an expression for the reaction quotient Q and (2) compare Q to K. **The expression for Q is the same as that for K, but the concentrations are not those at equilibrium**.

If Q<Ksp, no precipitate is formed.

If Q>Ksp, precipitate is formed.

Example: A solution is prepared by mixing 100.0 mL of 1.0 x 10-2 M $Pb(NO_3)_2$ and 100.0 mL of 1.0 x 10^{-3} M NaF. Will PbF_2(s) (K_{sp} = 4 x 10^{-8}) precipitate?

1. Start by writing the equation: $PbF_2(s) \rightleftharpoons Pb^{2+} + 2F^-$
2. Find <mark>initial concentrations</mark> of each ion. Be sure to account for dilution:
   a. $[Pb(NO_3)_2] = \frac{1.0*10^{-2}M*.100\ L}{.100\ L + .100\ L} = .005\ M = [Pb^{2+}]$
   b. $[NaF] = \frac{1.0*10^{-3}M*.100L}{.100\ L + .100\ L} = .0005\ M = [F^-]$
3. Write an <mark>expression for the reaction quotient (Q).</mark> Then, plug in the values obtained in (2).
   a. $Q = [Pb^{2+}][F^-]^2 = (.005)(.0005)^2 = 1.25 * 10^{-9}$
4. <mark>Compare Q to K.</mark>$1.25 * 10^{-9} < 4 * 10^{-8}$, so Q<K, so **no precipitate is formed**.

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Highlight 3: Selective Precipitation

Selective precipitation: If there is an ion in solution that is unwanted, it can be precipitated out, leaving other (wanted) ions behind. But it is necessary to make sure that the unwanted ion will precipitate **first**. You choose an ion that will form precipitates with the ions in solution, and as you add it, here is what happens:

Stage 1: Neither precipitate has begun forming (below solubility for both ions)

Stage 2: One combination of ions reaches saturation; a precipitate begins forming.

Stage 3: The other combination of ions reaches saturation as well; a second precipitate begins forming.

The trick is to choose ion combinations for which Stage 2 will precipitate out your unwanted ion. Then, stop adding ion after Stage 2 and before Stage 3. Neat trick, right? Don't worry too much about this unless your prof has shown practice problems!

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Check Your Understanding

1. What is the molar concentration of [Ag+] in AgCl solution in 0.10 M NaCl ? (Ksp = 1.8 x 10-10) Remember that in this case the molar solubility of AgCl is equal to the [Ag+] as only the Ag+ reflects the amount of AgCl that dissolved. (Problem 4 pg 2 from link)
2. Will mixture precipitate if equal volumes of 3.0x10-3M Ba+2 & 2.0x10-3M CO3 -2 mixed? Ksp for BaCO3 is 5.0x10-9 (Sample Problem 5 from link)

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Things You May Struggle With

1. Solubility is the amount of an ionic compound that can be added to solution before the solution begins rejecting it and forming a precipitate. Know the difference between this and solubility

product constant: Solubility product constant is the K value (review Equilibrium), and its expression involves solubility.

2. Different ions have different K expressions. These depend on the stoichiometry! It is a common mistake to assume that $K=x^2$.

3. Comparing Ksp values of different salts does not necessarily determine relative solubilities. Calculate the solubilities using the K expression, and compare those. If the salts have different stoichiometry, their K expressions will be different, and so the K values won't be comparable.

4. If a common ion is present, be sure to include its concentration in the "I" row.

5. If calculating equilibrium concentration after precipitation occurs: calculate how much excess ion is present. Then, use that as the "I" row of your I/C/E table. Don't use the full amount of ion added.

6. When calculating initial concentrations, if multiple solutions were mixed, be sure to include the effect of dilution.

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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 Learning are below.

1. 1.3*10^-5 M
2. Yes, Q=1.5*10^-6

ksp_problems_2.pdf (weebly.com)