5.4 - Special K's - Solubility Equilibrium

pages 545-553 in Health

Table 13 in resource package

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Since saturated solutions are equilibrium systems, we can apply the equilibrium constant expression to solutions.

Since a solution is a special kind of equilibrium, we will be dealing with a special kind of equilibrium constant, denoted K_{sp}. The sp stands for 'solubility product'.

Recall that in a solution, we often deal with aqueous ions. Therefore, dissociation will need to be used.

For example, consider a saturated silver sulfate solution:

$$Ag_2SO_{4(s)} \rightleftarrows 2Ag^+_{(aq)} + SO_4^{2-}{}_{(aq)}$$

Since solids and liquids are not involved in an K_{eq} calculation,

$$K_{eq} = [Ag^+]^2\,[SO_4^{2-}]$$

This equation only contains the products of a solution. Therefore, it is a special equilibrium constant that shows up a lot in chemistry.

$$K_{sp} = [Ag^+]^2\,[SO_4^{2-}]$$

Note that we **must** include the charges on the ions in this expression. Ag is different from Ag^+.

Note that unless otherwise indicated, K_{sp} calculations will always involve dissolving a solid to produce aqueous ions.

Also note that if you compare two K_{sp}'s together, the larger one will indicate a more soluble salt since the products will be favored. Therefore, a smaller K_{sp} indicates a less soluble salt because the reactants will be favored.

Ex) Write the solubility product constant expression for $Ca_3(PO_4)_2$.

Ex) The concentration of lead ions in a saturated solution of PbI_2 at 25°C is 1.3×10^{-3} M. What is the solubility product constant?

Ex) The K_{sp} for $MgCO_3$ at 25°C is 2.0×10^{-8}. What are the ion concentrations in a saturated solution at this temperature?

Ex) Calculate the solubility for silver chromate in a saturated solution at 25°C. $K_{sp} = 1.1 \times 10^{-12}$.

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Recall from Unit 4 that the *solubility* of a substance is the maximum concentration a substance can have in water to make a saturated solution.

Even 'insoluble' substances will dissolve slightly. However, the amount that actually dissolves has such a small concentration that we say that the substance is insoluble (less than 0.1M).

Also recall that we learned how to use a solubility table to determine whether or not a precipitate would form during a double displacement reaction. When doing those predictions, we **assumed** that the concentration of the 'insoluble' substance was great enough that it would accumulate at the bottom of the solution.

However, we could run into a situation where this 'insoluble' substance forms in such a small amount that the water is able to hold it. Therefore, no solid would form.

Remember that equilibrium systems involving solutions are **saturated**. Therefore, K_{sp} values are related to **saturated** solutions. If a solution is unsaturated then more substance can dissolve. If a solution is supersaturated then a solid should form at the bottom of the solution.

So, when we are trying to determine whether or not a precipitate will form when two ions meet each other, we need to determine the level of saturation. We can do this by comparing a **trial K_{sp}** to the **actual K_{sp}**.

The trial K_{sp} can be denoted as Q.

Q can be thought of as '**what we have**' and K_{sp} is '**what we need for saturation**'.

There are three situations that can happen when we compare Q to K_{sp}:

1. $Q < K_{sp}$

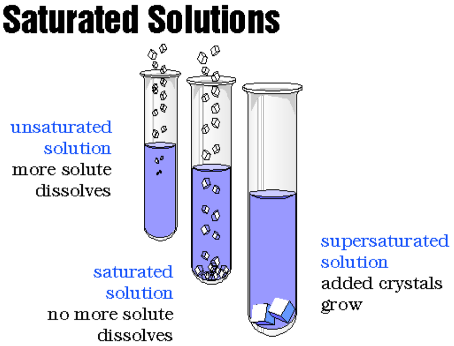
This means that 'what we have' is less than 'what we need'. Therefore, this represents an unsaturated solution and **no precipitate will form.**

2. $Q = K_{sp}$

This means 'what we have' is equal to 'what we need'. Therefore, this represents a saturated solution and **no precipitate will form.**

3. $Q > K_{sp}$

This means 'what we have' is more than 'what we need'. Therefore, this represents a supersaturated solution and **a precipitate will form.**



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What if I mix two substances together?.... they dilute each other.

This dilution must be considered when deciding on whether a precipitate will form (remember determining ion concentration from Unit 4...)

Steps:

1. Decide if any possible precipitates can form from the two substances being mixed using a solubility table.

   Write the net ionic equation for this substance to determine what the trial K_{sp} expression is.

2. Calculate the dilution of **both** substances present. Note that the final volume will be the sum of the volumes of both substances ($M_1V_1=M_2V_2$).

   Then, calculate the concentrations of the **ions** that make up your possible precipitate.

3. Using the concentrations of the **ions** above, calculate the trial K_{sp} and compare it to the actual K

Ex) 25.0 mL of 0.00200 M of potassium chromate are mixed with 75.0 mL of 0.000125 M of lead (II) nitrate. Will a precipitate form if K_{sp} of lead (II) chromate is 1.8×10^{-14}?

Ex) If 25.0 mL of 4.50×10^{-3} M $Pb(NO_3)_2$ is mixed with 35.0 mL of 2.80×10^{-3} M MgI_2, will a precipitate form?

-We can also use the K_{sp} to determine the maximum concentration of an ion which can exist in solution with another ion without precipitation

Ex) Water hardness is caused by the presence of Ca^{2+} and Mg^{2+} ions. One way of removing these ions is to add washing soda (sodium carbonate Na_2CO_3) which causes the precipitation of $CaCO_3$ and $MgCO_3$. If 5.0L of water has $[Ca^{2+}]$ of 0.0040M, calculate the maximum mass of Na_2CO_3 which can be added without causing any precipitate to form. K_{sp} for $CaCO_3$ = 4.8 x 10-9

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5.4 – Special K's – Equilibrium Solubility - Assignment

1. Write the balanced equation and the solubility product constant expression, K_{sp}, for the each of the following dissociation reactions. All compounds are solids. One has been given as an example.

Reminders – ion charges MUST BE included.

– solids (and liquids) are NOT included in the equilibrium expression

– don't forget to include exponents when needed

– polyatomic ions (e.g. CO_3^-) do not break apart

Compound	Equation	Ksp
$(NH_4)_2S$	$(NH_4)_2O\ (s) \rightleftharpoons 2\ NH_4^+(aq) + S^{2-}(aq)$	$K_{sp} = [NH_4^+]^2[S^{2-}]$
CaS		
K_2SO_4		
$Mg(OH)_2$		

2. Organize the following salts in order of solubility (highest to lowest):

AgCl; $K_{sp} = 1.8 \times 10^{-10}$ AgI; $K_{sp} = 8.5 \times 10^{-17}$ AgBr; $K_{sp} = 5.4 \times 10^{-13}$

3. Calculate K_{sp} for a saturated nickel(II) sulfide, NiS, solution with a solubility of 3.27×10^{-11}. Calculate the K_{sp}.

4. Calculate the concentration of ions in a saturated solution of $CaCO_3$ in water at 25°C. K_{sp} for $CaCO_3$ is 4.8×10^{-9}.

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5. Calculate the concentrations of ions at 25°C for a saturated solution of silver bromate (Ksp = 5.3×10^{-5}).

6. At 25°C, 0.0024 g of $Ce(OH)_3$ is contained in a 2.5 L solution. Calculate K_{sp}.

7. What is the mass of calcium fluoride present in a saturated 1.5 L solution?

8. 400.0 mL of 4.00×10^{-10} M $Al(NO_3)_3$ is mixed with 500.0 mL of 3.00×10^{-7} M NaOH. If K_{sp} for $Al(OH)_3$ is 5.00×10^{-33} at this temperature, will there be a precipitate?

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9. Will a precipitate form if 20.0 mL of 0.0100 M $CaCl_2$ are mixed with 20.0 mL of 0.00800 M Na_2SO_4 at 25.0 ˚

10. Will a precipitate form if 25 mL of 4.0 x 10^{-3} M $AgNO_3$ are mixed with 75 mL of 2.0 x 10^{-4} M Na_2CrO_4 at 25.0 ˚C?

11. What is the maximum [Sr^{2+}] that can be dissolved in a 0.020 M solution of K_2SO_4 without precipitating $SrSO_4$? (K_{sp} of $SrSO_4$ = 7.6 x 10^{-7})