7.6: SOLUBILITY AND SOLUBILITY PRODUCT

Solubility is the maximum amount of the substance, which will dissolve at a given temperature (at this point the solution is saturated)
A saturate solution is a solution in which dynamic equilibrium exists between undissolved and the dissolved solute

- Unsaturated Solution??
- Super Saturated solution??

A substance is considered soluble if it will form a solution of concentration.> 0.1 moles per litre at room temperature otherwise it is insoluble.

GENERAL SOLUBILITY RULES
A low value of Ksp means the concentrations of ions are low at equilibrium. Hence the solubility must be low.

SOLUBILITY PRODUCT
When a solid electrolyte is in a state of equilibrium with a saturated solution the concentration of a solid is a constant.
The equilibrium constant is called **SOLUBILITY PRODUCT CONSTANT**
This new constant, equal to the products of the concentrations of all the ions present is called SOLUBILITY PRODUCT CONSTANT **(Ksp)**.

CALCULATING Ksp FROM SOLUBILITY

Example1: Copper (1) Bromide has a measured solubility of 2.0×10^{-4} mol/L at $25^{o}C$. Calculate its Ksp value

Solution

$CuBr\ (s) \leftrightarrow Cu^{+}_{(aq)} + Br^{-}_{(aq)}$

$Ksp = [Cu^{+}][Br^{-}]$

$[CuBr] = [Cu^{+}] = [Br^{-}]$
$[Cu+] = [Br-]$

$Ksp = (2.0 \times 10^{-4})(2.0 \times 10^{-4})$

$= 4.0 \times 10^{-8}$ (There is no unit)

Or Set an ICE Table As Follows:

	$CuBr\ (s) \leftrightarrow$	$Cu^{+}_{(aq)}$	$+\ Br^{-}_{(aq)}$
Initial conc.	-	0	0
Change conc.	-	x	x
Equil Conc.	-	x	x

$x = 2.0 \times 10^{-4}$
$x = 2.0 \times 10^{-4}$

$Ksp = [Cu+ (aq)][Br- (aq)]$

$=(2.0 \times 10^{-4})(2.0 \times 10^{-4})$
$= 4.0 \times 10^{-8}$

Example 2: Calculate the Ksp value for bismuth sulfide ($Bi_2S_{3(s)}$). Which has solubility of 1.0×10^{-15} mol/L at 25 ^{o}C.

Solution:

$Bi_2S_{3(s)} \leftrightarrow [2Bi^{+3}\ (aq)][3S^{-2}\ _{(aq)}]$

$Ksp = [Bi^{3+}]2[S^{-2}]^3$.

Since no Bi^{3+} and S^{2-} ions were present in solution before the Bi2S3 dissolved,
$[Bi^{3+}]_0 = [S^{2-}]_0 = 0$
Thus the equilibrium concentration is determined by the amount of the salt that dissolved to reach equilibrium, which in this case is 1.0×10^{-15} mol/L

$Ksp = 2[Bi^{3+}]^2 3[S^{2-}]^3$

$=2(1.0 \times 10^{-15})^2 3(1.0 \times 10^{-15})^3$

$= (2.0 \times 10^{-15})^2(3.0 \times 1.0^{-15})^3$

$= 1.1 \times 10^{-73}$

OR Set an ICE table

	$Bi_2S_{3\,(s)} \leftrightarrow$	$2Bi^{3+}_{(aq)}$ +	$3S^{2-}_{(aq)}$
Initial conc.	-	0	0
Change conc.	-	2x	3 x
Equil Conc.	-	2x	3x

$x = 1.0 \times 10^{-15}$
$2x = 2(1.0 \times 10^{-15})$
$= 2.0 \times 10^{-15}$
$3x = 3(1.0 \times 10^{-15})$
3.0×1.0^{-15}
$Ksp = [Bi^{3+}_{(aq)}]^2[S^{2-}_{(aq)}]^3$

$=(2.0 \times 10^{-15})^2(3.0 \times 10^{-15})^3$

$= (4.0 \times 10-30)(2.7 \times 10-44)$

$= 1.1 \times 10^{-73}$

Calculating Solubility from Ksp
When an ionic substance dissolves in water to form a saturated solution, equilibrium is created between the solid and its constituent ions.

- The concentration of the saturated solution is known as the SOLUBILITY OF THE SUBSTANCE

Example 1:
Calculate the solubility of Zinc hydroxide at 25°C. The Ksp of Zinc hydroxide(s) is 4.5×10^{-17} at 25°C

Solution:

$Zn(OH)_{2(s)} \leftrightarrow Zn^{2+}_{(aq)} + 2(OH)^-_{(aq)}$

$[Zn(OH)_{2(aq)}] = [Zn^{2+}_{(aq)}]$

The equilibrium Expression is

$Ksp = [Zn^{2+}_{(aq)}][OH_{(aq)}]^2 = 4.5 \times 10^{-17}$

ICE Table

	$Zn(OH)_{2(s)} \leftrightarrow$	$Zn^{2+}_{(aq)}$ +	$2(OH)^-_{(aq)}$
Initial Conc	-	0	0

Change Conc	-	x	2x
Equilib Conc.	-	x	2x

$Ksp = [Zn^{2+}_{(aq)}][OH_{(aq)}]^2 = 4.5 \times 10^{-17}$

$[(OH)^-_{(aq)}] = 2[Zn^{2+}_{(aq)}]$

$Ksp = [Zn^{2+}_{(aq)}] (2[Zn^{2+}_{(aq)}])^2$

Ksp = (x)(4x2)

4.5×10^{-17} = (x)(4x2)

4.5×10^{-17} = (4x3)

$x^3 = \frac{4.5 \times 10^{-17}}{4}$

$\mathbf{x} = \sqrt[3]{\frac{4.5 \times 10^{-17}}{4}}$

$[Zn^{2+}_{(aq)}] = 2.24 \times 10^{-6}$

$[Zn(OH)_{2(aq)}] = [Zn^{2+}_{(aq)}] = 2.24 \times 10^{-6}$

Practice P. 488 # 1,2 & 4

SOLUBILITY RULES

Potassium, sodium and ammonia salts
Whatever they may be,
Can always be relied upon
for solubility.
Every single sulfate
Is soluble 'tis said,
"Cept barium and calcium
and strontium and lead.
Most every chloride's soluble.
That's what we've always read.
Save silver, mercurous mercury,
And (slightly) chloride of lead.
When asked about the nitrates,
The answer's always clear,
They each and all are soluble,
That's all we want to hear.
Metallic bases won't dissolve,
That is, all but three:
Potassium, sodium and ammonium
Dissolve quite readily.
But then you must remember
That you must not forget
Calcium and barium
Dissolve a little bit.
Carbonates are insoluble,
It's lucky that it's so,
Or all our marble buildings
Would melt away like snow.

Predicting Precipitation

Recall the use of the value of reaction quotient, Q, to determine the direction of equilibrium in a system.
It can also be used to predict whether a precipitate will form or not when we mix solutions.
The reaction quotient is called a trial ion product in this case. We compare the Ksp of the salt of these ions to Q (trial ion product).
If Q is higher than Ksp, a precipitate will form, and if Q is smaller than Ksp, a precipitate will not form. Also, if Q is equal to Ksp, a precipitate will not form.

Example:
Predict whether a precipitate will form if 25.0 mL of 0.010 mol/L silver nitrate is mixed with 25.0 mL of 0.0050 mol/L potassium chloride
$AgNO_{3(aq)} + KCl_{(aq)} \rightarrow AgCl_{(s)} + KNO_{3(aq)}$

This is a double displacement reaction. From above, AgI(s) is relatively insoluble. But are the concentration of Ag+ and I- high enough to form a precipitate?
$AgNO_{3(aq)} \leftrightarrow Ag^{+}_{(aq)} + NO_{3(aq)}$

$[Ag^{+}_{(aq)}] = [NO_3^{-}{}_{(aq)}] = 0.010$ mol/L (before mixing)

$KCl_{(aq)} \leftrightarrow K^{+}_{(aq)} + Cl^{-}_{(aq)}$

$[Cl^{-}_{(aq)}] = [KCl\ (aq)] = 0.0050$ mol/L (before mixing)

$$[Ag^{+}_{(aq)}] = 0.010 \text{ mol/L} \times \frac{25 \text{ mL}}{50.0 \text{ mL}} = 5 \times 10^{-3} \text{ mol/L}$$

$$[Cl^{-}_{(aq)}] = 0.0050 \text{ mol/L} \times \frac{25 \text{ mL}}{50.0 \text{ mL}} = 2.5 \times 10^{-3} \text{ mol/L}$$

Then Calculate Q for AgCl(s) $\quad K_{sp} = 1.8 \times 10^{-10}$
$AgCl(s) \leftrightarrow Ag+ (aq) + Cl- (aq)$

$Q = [Ag^{+}_{(aq)}][Cl^{-}_{(aq)}]$
$(5 \times 10^{-3})(2.5 \times 10^{-3}) = 1.25 \times 10^{-5}$
Since Q is higher Ksp, a precipitate will form.

Practice P. 489 # 5b,c
Make notes on Common ion Effect P.490 Q # 1-11 P. 493