Introduction to solubility equilibrium

High solubility - no equilibrium is set up the solution is aqueous

Low solubility - a precipitate forms and an equilibrium occurs

An equilibrium occurs between the precipitate and the dissolved ions

The equilibrium constant is a product of the two concentrations of dissolved ions raised to the power of their coefficient in a dissociation equation

Ex. $AgCl_{(s)} \leftrightarrow Ag^{+}_{(aq)} + Cl^{-}_{(aq)}$

$$K = [Ag^{+}][Cl^{-}]$$

When setting up an ICE table the x value corresponds to the solubility of the chemical. The solubility of the chemical must be in moles per liter

	$AgCl_{(s)}$ $\leftrightarrow$	$Ag^{+}_{(aq)}$ +	$Cl^{-}_{(aq)}$
I	some amount	0	0
C	-x	+x	+x
E	(solid not included)	x	x

When given the equilibrium constant you can solve for the solubility

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Trial Ion product

Like a Q value the trial ion product value allows you to predict the direction of the equilibrium equation (dissociation).

For solubility it allows us to predict if a precipitate will form

1. Calculate the concentrations of the ions that form a precipitate
2. Substitute them into the equilibrium expression
3. Calculate the Q value
4. Compare the Q value to the equilibrium constant

- If the Q value > the equilibrium constant a precipitate forms
- If the Q value < the equilibrium constant no precipitate forms
- If the Q value = equilibrium constant a saturated solution has occurred

Common ion effect

The presence of an ion that may participate in the formation of a precipitate will alter the solubility of the chemical that forms a precipitate

In developing the ICE table the initial concentration of an ion is no longer zero

Continue to solve for the x value which will be the new solubility of that chemical