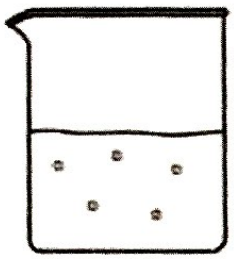
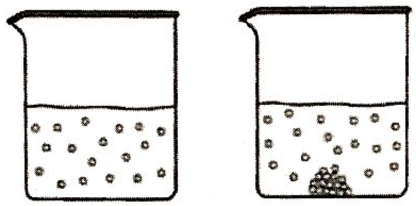
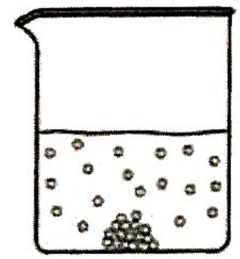
Will a Precipitate Form? A Task for K vs Q!

Precipitation occurs when the concentrations of ions is greater than the solubility of the ionic compound.

Compare the value of Q with given K_{sp} to determine if a precipitate will form!

K > Q	**K = Q**	**K < Q**
Unsaturated Solution	Saturated Solution	Saturated Solution with extra
System will shift right to reach equilibrium $AgCl(s) \rightleftarrows Ag^+(aq) + Cl^-(aq)$	At equilibrium $AgCl(s) \rightleftarrows Ag^+(aq) + Cl^-(aq)$	System will shift left to reach equilibrium $AgCl(s) \rightleftarrows Ag^+(aq) + Cl^-(aq)$
More solid will dissolve until K = Q.	Solid will both dissolve and precipitate at the same rate.	**More solid will precipitate** until K = Q.
 no precipitate	 maybe precipitate	 yes precipitate

Important Ideas to Note:

1. If any solid is present, after a period of time the solution is at equilibrium (a Saturated solution)
2. Ion concentration, $\boxed{\text{[ions], is } \textbf{independent} \text{ of volume when at equilibrium}}$ (for instance, in a Saturated solution).
3. If ions are present that could form multiple salts, the solid with the smallest molar solubility will form.

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Solubility Equilibrium Translation Guide

1. Solubility product constant = K_{sp} (aka the equilibrium constant for solubility)
2. Molar solubility = x from RICE table (aka how many moles of a solid will dissolve in 1.0 L, units = M = mol/L)
3. Saturated = equilibrium (aka a solution has dissolved as many ions as can fit, any extra will precipitate)

Let's Practice!

1. A chemist makes a 2.0 L saturated solution of $Ba_3(PO_4)_2$ solution, which has a $K_{sp} = 6.0 \times 10^{-39}$.

   a. What is the concentration of Ba^{2+} ions in solution?

$$K_{sp} = [Ba^{2+}]^3[PO_4^{2-}]^2 = (3x)^3(2x)^2 = 108x^5 = 6.0\text{E-}39$$

$$\Rightarrow x = \sqrt[5]{\frac{6.0\text{E-}39}{108}} = 8.9\text{E-}9 \text{ M}$$

$$[Ba^{2+}] = 3x = 3(8.9\text{E-}9) = \boxed{2.7\times10^{-8} \text{ M}}$$

b. After two days of sitting on the counter, some liquid has evaporated from the solution. Did $[Ba^{2+}]$ increase, decrease, or remain the same? Justify your answer.

Remain the same! Although less liquid is present, the sol'n was already saturated, thus $[Ba^{2+}]$ cannot increase.

c. The chemist adds 3.00 g of solid $(NH_4)_3PO_4$ to the original saturated solution of $Ba_3(PO_4)_2$. Did $[Ba^{2+}]$ increase, decrease, or remain the same? Justify your answer.

Decrease, b/c adding a common ion, PO_4^{3-}, to sol'n will increase [products], + so the rxn will shift left to re-establish equilibrium. This means more $Ba_3(PO_4)_2$ (s) will form, removing some Ba^{2+} from sol'n.

2. A solution containing lead (II) nitrate is mixed with one containing sodium bromide to form a solution that is 0.0150 M in $Pb(NO_3)_3$ and 0.00350 M NaBr. Does a precipitate form in this newly mixed solution? (K_{sp} of $PbBr_2$ = 4.67×10^{-6})

$$Q = [Pb^{2+}][Br^-]^2 = (0.0150)(0.00350)^2 = 1.84 \times 10^{-7}$$

$K > Q$ $K > Q$ so no precipitate will form.

$4.67E\text{-}6 > 1.84E\text{-}7$

3. The K_{sp} value for lead (II) bromide, $PbBr_2$, is 4.6×10^{-6} at 25°C. What is the maximum mass, in grams, of $PbBr_2$ that can dissolve in 1.50 L of water?

$$K_{sp} = [Pb^{2+}][Br^-]^2 = x(2x)^2 = 4x^3 = 4.6E\text{-}6$$

$$x = \sqrt[3]{\frac{4.6E\text{-}6}{4}} = 0.01048 \text{ M } PbBr_2 \times 1.50 \text{ L} = 0.016 \text{ mol} \times \frac{367 \text{ g } PbBr_2}{1 \text{ mol}}$$

$$= \boxed{5.8 \text{ g } PbBr_2}$$