Predicting a Precipitate

Recall:

$$\mathbf{AB\ (s) \rightleftharpoons A^{+}\ (aq) + B^{-}\ (aq)}$$

In a saturated solution at equilibrium, $K_{sp} = [A^+]\,[B^-]$

What if $[A^+][B^-] < K_{sp}$?

There would be no precipitate (solid) formed in the beaker.

What if $[A^+][B^-] > K_{sp}$?

If this occurs, a solution is supersaturated. A precipitate will form:

[A+] ↓ and [B−] ↓ until their product $[A^+][B^-]$ = Ksp.

Reaction Quotient and Formation of a Precipitate

Based on the solubility rules, calcium carbonate ($CaCO_3$) is insoluble and should form a precipitate. However, tap water contains low concentrations of both these ions. Why doesn't tap water contain a precipitate?

This is because their concentrations are too low to form a solid.

$CaCO_{3\,(s)} \rightleftharpoons Ca^{2+}{}_{(aq)} + CO_3^{2-}{}_{(aq)}$ $\quad K_{sp} = [Ca^{2+}][CO_3^{2-}] = 3.36 \times 10^{-9}$

If the product of $[Ca^{2+}][CO_3^{2-}]$ is *less* than the K_{sp}, no precipitate forms.

Reaction Quotient (Q) or Trial Ion Product (TIP):

Takes on the same form as the equation for K_{sp} but uses the actual concentrations at that time. Once Q or TIP is calculated it can be compared to K_{sp} to see if a precipitate forms.

If:	$Q < K_{sp}$	the system is unsaturated - no precipitate forms
	$Q = K_{sp}$	the system is saturated - no precipitate forms
	$Q > K_{sp}$	the system is supersaturated - a precipitate forms

Quantitative Determination of Precipitation

1) If 1×10^{-6} mol of solid $AgNO_3$ and 5×10^{-4} mol of solid Na_2S are added to 10 L of water, will a precipitate form? K_{sp} (Ag_2S) = 8.8×10^{-18}

2) If 20 mL of 0.005 mol/L barium nitrate is added to 80 mL of 0.0001 mol/L sodium carbonate, will a precipitate form? K_{sp} ($BaCO_3$) = 8.1×10^{-9}

3) Will a precipitate form if 0.25 mol of Na_2CrO_4 is added to 5.0 L of a 0.020 mol/L solution of $AgNO_3$? K_{sp} (Ag_2CrO_4) = 8.0×10^{-4}