Will a Precipitate Actually Occur? Using TIPs

We can predict the products of a double displacement reaction to determine if a precipitate is likely to form. However, this precipitate will only be created if there are enough ions in solution to establish equilibrium.

As such, we need to do the math to determine if we have enough ions! Called the TRIAL ION PRODUCT (TIP), this math is basically a trial Keq / Ksp.

The end result of TIP is symbolized using a Q, which stands for the product of the ion concentrations which actually exist in solution.

Long story short: calculate the TIP using the values given in the problem (assuming they are equilibrium values) and then compare TIP to K_{sp}.

If: $Q < Ksp$ Not enough ions to form a precipitate
$Q = Ksp$ Adequate ions to form a saturated solution but not enough for a precipitate
$Q > Ksp$ Excess of ions and a precipitate will form

Steps:
1. Balanced Double Displacement Reaction
2. Net Ionic Equation for the Precipitate
3. Find Concentration of Ions $\quad Concentration\ x\ \frac{Volume}{Total\ Volume}$
4. Determine TIP equation / Substitute Concentrations into "Q"
5. Compare Q to Ksp

<u>Example:</u> If the K_{sp} for $PbCl_2$ is 1.8×10^{-4}, will a precipitate form when 200.0 mL of 0.015 M NaCl is mixed with 100.0 mL of 0.060 M $Pb(NO_3)_2$?

(1) $2NaCl_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow 2NaNO_{3(aq)} + PbCl_{2(s)}$

(2) $Pb^{+2}_{(aq)} + 2Cl^-_{(aq)} \rightarrow PbCl_{2(s)}$

(3) $[Pb^{+2}] = 0.060\ M \times \left(\frac{100mL}{300mL}\right) = 0.02\ M\ [Pb^{+2}]$

$[Cl^-] = 0.015\ M \times \left(\frac{200mL}{300mL}\right) = 0.01\ M\ [Cl^-]$

(4) $Q = [Pb^{+2}][Cl^-]^2$
$= [0.02][0.01]^2$
$= 2 \times 10^{-6}$

(5) $Ksp = 1.8 \times 10^{-4}$ while $Q = 2.0 \times 10^{-6}$

$\boxed{Q < ksp \therefore \text{no precipitate will form}}$

Example: Will a precipitate form if 25.0 mL of 1.4×10^{-9} mol L^{-1} sodium iodide and 35.0 mL of 7.9×10^{-7} mol L^{-1} silver nitrate are mixed? The K_{sp} for AgI at 25°C is 8.5×10^{-17}.

$NaI_{(aq)} + AgNO_{3(aq)} \rightarrow AgI_{(s)} + NaNO_{3(aq)}$

$Ag^{+}_{(aq)} + I^{-}_{(aq)} \rightarrow AgI_{(s)}$

$[Ag^+] = 7.9 \times 10^{-7} M \times \left(\frac{35 mL}{60 mL}\right) = 4.61 \times 10^{-7} M \; [Ag^+]$

$[I^-] = 1.4 \times 10^{-9} M \times \left(\frac{25 mL}{60 mL}\right) = 5.83 \times 10^{-10} M \; [I^-]$

$Q = [Ag^+][I^-]$

$= [4.61 \times 10^{-7}][5.83 \times 10^{-10}]$

$= 2.7 \times 10^{-16}$

$K_{sp} = 8.5 \times 10^{-17}$

$\boxed{Q > K_{sp} \therefore \text{ a precipitate will form}}$

Example: A student prepares a solution containing 0.010M calcium nitrate and 0.025 M sodium phosphate. Will a precipitate form when the student makes this solution? If so, identify the precipitate.

$3\,Ca(NO_3)_{2(aq)} + 2\,Na_3PO_{4(aq)} \rightarrow 6\,NaNO_{3(aq)} + Ca_3(PO_4)_{2(s)}$

$3\,Ca^{+2}_{(aq)} + 2\,PO_4^{-3}{}_{(aq)} \rightarrow Ca_3(PO_4)_{2(s)}$

$[Ca^{+2}] = 0.01 M$ from $Ca(NO_3)_2$ 1:1 $\therefore [Ca^{+2}] = 0.01 M$

$[PO_4^{-3}] = 0.025 M$ from Na_3PO_4 1:1 $\therefore [PO_4^{-3}] = 0.025 M$

$Q = [Ca^{+2}]^3[PO_4^{-3}]^2$

$= [0.01]^3[0.025]^2$

$= 6.25 \times 10^{-10}$

$K_{sp} = ?$ Look up! p. 725 in text

2.1×10^{-33}

$\boxed{Q > K_{sp} \therefore \text{ yes a precipitate will form}}$