	Marks
• A mixture of NaCl (5.0 g) and $AgNO_3$ (5.0 g) was added to 1.0 L of water. What are the concentrations of Ag^+(aq), $Cl^–$(aq) and Na^+(aq) ions in solution after equilibrium has been established? K_{sp}(AgCl) = 1.8 × $10^{–10}$.	**3**

The molar masses of the two salts are:

$$\boldsymbol{M_{NaCl} = (22.99\ (Na) + 35.45\ (Cl))\ g\ mol^{-1} = 58.44\ g\ mol^{-1}}$$

$$\boldsymbol{M_{AgNO_3} = (107.87\ (Ag) + 14.01\ (N) + 3 \times 16.00)\ g\ mol^{-1} = 169.88\ g\ mol^{-1}}$$

The number of moles of salt added to the solution are therefore:

$$\textbf{number of moles of NaCl} = \frac{\textbf{mass}}{\textbf{molar mass}} = \frac{\mathbf{5.0}}{\mathbf{58.44\ g\ mol^{-1}}} = \textbf{0.0855 mol}$$

$$\textbf{number of moles of } \mathbf{AgNO_3} = \frac{\mathbf{5.0}}{\mathbf{169.88\ g\ mol^{-1}}} = \textbf{0.0294 mol}$$

As 1.0 L of water is present, the initial concentrations of the ions are [Na^+(aq)] = 0.086 M, [Cl^-(aq)] = 0.086 M and [Ag^+(aq)] = 0.029 mol. The Na^+(aq) will form any precipitate with the ions present: [Na^+(aq)] = 0.086 M.

The ionic product for the precipitation of AgCl(s) is given by:

$$\boldsymbol{Q_{sp} = [Ag^+(aq)][Cl^-(aq)] = (0.029)(0.086) = 0.0025}$$

As Q_{sp} >> K_{sp}, precipitation of AgCl(s) will occur. As [Ag^+(aq)] < [Cl^-(aq)], the silver ion concentration is limiting and so:

$$\boldsymbol{[Cl^-(aq)] = (0.086 – 0.029)\ M = 0.056\ M}$$

As AgCl(s) is present, [Ag^+(aq)] is given by the solubility product:

$$\boldsymbol{K_{sp} = [Ag^+(aq)][Cl^-(aq)] = 1.8 \times 10^{-10}}$$

$$\boldsymbol{[Ag^+(aq)] = (1.8 \times 10^{-10}) / (0.056)\ M = 3.2 \times 10^{-9}\ M}$$

[Ag^+(aq)] = **3.2 × 10^{-9} M**	[$Cl^–$(aq)] = **0.056 M**	[Na^+(aq)] = **0.086 M**