122. Magnesium fluoride, $MgF_2(aq)$ has a molar solubility of 2.7×10^{-3} mol/L. Use this information to determine the K_{sp} value for the solid.

What Is Required?
You need to calculate the K_{sp} for magnesium fluoride, MgF_2.

What Is Given?
You know the solubility of magnesium fluoride is 2.7×10^{-3} mol/L.

Plan Your Strategy	Act on Your Strategy
Write a balanced equation for the solubility equilibrium.	$MgF_2(s) \rightleftharpoons Mg^{2+}(aq) + 2F^-(aq)$
Write the expression for the K_{sp}.	$K_{sp} = [Mg^{2+}][F^-]^2$
Determine the concentration of each ion. 1 mol of $MgF_2(s)$ ionizes to give 1 mol of Mg^{2+} and 2 mol F^- ion.	$[Mg^{2+}] = [MgF_2] = 2.7 \times 10^{-3}$ mol/L $[F^-] = 2[MgF_2]$ $= 2 \times 2.7 \times 10^{-3}$ mol/L $= 5.4 \times 10^{-6}$ mol/L
Substitute the ion concentrations into the K_{sp} expression and solve.	$K_{sp} = [Mg^{2+}][F^-]^2$ $= (2.7 \times 10^{-5})(5.4 \times 10^{-5})^2$ $= 7.9 \times 10^{-8}$

Check Your Solution
The calculated value of K_{sp} has the correct number of significant digits that agrees with the given data. The small value is expected for a compound of low solubility.

123. Silver sulfide, $Ag_2S(aq)$ has a K_{sp} value that is equal to 5.6×10^{-49}. What is the molar solubility of the solid?

What Is Required?
You need to calculate the molar solubility of silver sulfide, $Ag_2S(s)$.

What Is Given?
You know the K_{sp} for silver sulfide is 5.6×10^{-49}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the solubility equilibrium.	$Ag_2S(s) \rightleftharpoons 2Ag^+(aq) + S^{2-}(aq)$
Write the solubility product equation. Let x represent the molar solubility of $Ag_2S(s)$.	$K_{sp} = [Ag^+]^2[S^{2-}]$
Summarize the solubility of the $Ag^+(aq)$ and $S^{2-}(aq)$ ions at equilibrium in an ICE table.	See the ICE table below.

	$Ag_2S(s)$	$\rightleftharpoons$ $2Ag^+(aq)$	+ $S^{2-}(aq)$
		$[Ag^+]$ (mol/L)	$[S^{2-}]$ (mol/L)
I			
C	$-x$	$+2x$	$+x$
E		$2x$	x

Write the expression for the K_{sp} and solve for x.	$K_{sp} = [Ag^+]^2[S^{2-}]$ $5.6 \times 10^{-49} = (2x)^2(x)$ $= 4x^3$ $x = 5.2 \times 10^{-17}$ mol/L

Check Your Solution
The solubility is low as expected for a compound with the given K_{sp}. The answer correctly shows 2 significant digits.

124. The K_{sp} value for lead(II) bromide, $PbBr_2(aq)$, is 6.6×10^{-6}. What is the solubility of the solid?

What Is Required?
You need to calculate the solubility of lead(II) bromide(s) in g/L.

What Is Given?
You know the K_{sp} for lead(II) bromide is 6.6×10^{-6}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the solubility equilibrium.	$PbBr_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Br^-(aq)$
Write the solubility product equation. Let x represent the molar solubility of $PbBr_2(s)$.	$K_{sp} = [Pb^{2+}][Br^-]^2$
Summarize the solubility of the $Pb^{2+}(aq)$ and $Br^-(aq)$ ions at equilibrium in an ICE table.	See the ICE table below.

	$PbBr_2(s)$	$\rightleftharpoons$ $Pb^{2+}(aq)$	+ $2Br^-(aq)$
		$[Pb^{2+}]$ (mol/L)	$[Br^-]$ (mol/L)
I			
C	$-x$	$+x$	$+2x$
E		x	$2x$

Write the expression for the K_{sp} expression and solve for x.	$K_{sp} = [Pb^{2+}][Br^-]^2$
Determine the molar mass, M, of $PbBr_2$ and express the solubility in g/L ($m = nM$).	$6.6 \times 10^{-6} = x(2x)^2$ $= 4x^3$ $x = 1.18 \times 10^{-2}$ mol/L $= 1.18 \times 10^{-2}$ ~~mol~~/L $\times$ 367.00 g/~~mol~~ $= 4.3$ g/L

Check Your Solution
The solubility is low as expected for a compound with the given K_{sp}. The answer correctly shows 2 significant digits.

125. Which solid can have more mass ionize into 1.00 L of solution: silver chloride, AgCl(s), or copper(I) chloride, CuCl(s)?

What Is Required?
You need to compare the solubility of silver chloride and copper(I) chloride in g/L.

What Is Given?
From Appendix B, the K_{sp} for silver chloride is 1.77×10^{-10} and for copper(I) chloride is1.72×10^{-9}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the solubility equilibrium of silver chloride.	$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$
Write the solubility product equation. Let x represent the molar solubility of AgCl(s).	$K_{sp} = [Ag^+][Cl^-]$
Write the expression for the K_{sp}. Let x represent the concentrations of Ag^+ and Cl^-. Solve for x. Determine the molar mass, M, of AgCl and express the solubility in g/L ($m = nM$).	$K_{sp} = [Ag^+][Cl^-]$ $1.77 \times 10^{-10} = (x)(x)$ $x^2 = 1.77 \times 10^{-10}$ $x = 1.33 \times 10^{-5}$ mol/L $= 1.33 \times 10^{-5}\ \cancel{mol}/L \times 143.32\ g/\cancel{mol}$ $= 1.91 \times 10^{-3}$ g/L
Repeat these steps for CuCl(s) and determine which will have the greater mass dissolved in 1 L of solution.	$CuCl(s) \rightleftharpoons Cu^+(aq) + Cl^-(aq)$ $K_{sp} = [Cu^+][Cl^-]$ $1.72 \times 10^{-9} = (x)(x)$ $x = 4.15 \times 10^{-5}$ mol/L $= 4.15 \times 10^{-5}\ \cancel{mol}/L \times 99.00\ g/\cancel{mol}$ $= 4.11 \times 10^{-3}$ g/L More CuCl(s) will dissolve per litre of solution.

Check Your Solution
The solubilities are expressed correctly to 3 significant digits and the answer is reasonable.

126. The solubility of nickel(II) phosphate, $Ni_3(PO_4)_2(aq)$, is 7.8×10^{-5} g/L. Determine the solubility-product constant for this solid.

What Is Required?
You need to calculate the K_{sp} for nickel(II) phosphate, $Ni_3(PO_4)_2(aq)$.

What Is Given?
You know the solubility of $Ni_3(PO_4)_2(aq)$ is 7.8×10^{-5} g/L.

Plan Your Strategy	Act on Your Strategy
Use the molar mass, M, of $Ni_3(PO_4)_2$ and the formula $n = \frac{m}{M}$ to express the solubility in mol/L.	$n = \frac{m}{M}$ $= \frac{7.8 \times 10^{-5}\ \cancel{g}/L}{366.01\ \cancel{g}/mol}$ $= 2.13 \times 10^{-7}$ mol/L
Write a balanced chemical equation for the solubility equilibrium.	$Ni_3(PO_4)_2(s) \rightleftharpoons 3Ni^{2+}(aq) + 2PO_4^{3-}(aq)$
Determine the concentration of each ion. 1 mol of $Ni_3(PO_4)_2$ ionizes to give 3 mol of $Ni^{2+}(aq)$ ions and 2 mol $PO_4^{3-}(aq)$ ions.	$[Ni^{2+}] = 3[Ni_3(PO_4)_2]$ $= 3 \times 2.13 \times 10^{-7}$ mol/L $= 6.39 \times 10^{-7}$ mol/L $[PO_4^{3-}] = 2[Ni_3(PO_4)_2]$ $= 2 \times 2.13 \times 10^{-7}$ mol/L $= 4.26 \times 10^{-7}$ mol/L
Write the expression for the K_{sp}.	$K_{sp} = [Ni^{2+}]^3[PO_4^{3-}]^2$
Substitute the ion concentrations into the K_{sp} expression and solve.	$K_{sp} = [Ni^{2+}]^3[PO_4^{3-}]^2$ $= (6.39 \times 10^{-7})^3(4.26 \times 10^{-7})^2$ $= 4.74 \times 10^{-32}$

Check Your Solution
The calculated value of K_{sp} has the correct number of significant digits that agrees with the given data. The small value for K_{sp} is expected for a compound of low solubility.

127. The solubility of strontium fluoride, $SrF_2(aq)$, is 12.2 mg/100 mL. What is K_{sp} for this solid?

What Is Required?
You need to calculate the K_{sp} for strontium fluoride, $SrF_2(s)$.

What Is Given?
You know the solubility of strontium fluoride, $SrF_2(aq)$, is 12.2 mg/100 mL.

Plan Your Strategy	Act on Your Strategy
Use the molar mass, M, of SrF_2 and the formula $n = \frac{m}{M}$ to express the solubility in mol/L.	$\frac{12.2\ \cancel{mg} \times \frac{1\ g}{1000\ \cancel{mg}}}{100\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}}} = 0.122\ g/L$ molar solubility $= \frac{0.122\ \cancel{g}/L}{125.62\ \cancel{g}/mol}$ $= 9.71 \times 10^{-4}$ mol/L
Write a balanced equation for the solubility equilibrium.	$SrF_2(s) \rightleftharpoons Sr^{2+}(aq) + 2F^-(aq)$
Determine the concentration of each ion. One mol of SrF_2 ionizes to give one mol of $Sr^{2+}(aq)$ ion and two mol $F^-(aq)$ ions.	$[Sr^{2+}] = [SrF_2]$ $= 9.71 \times 10^{-4}$ mol/L $[F^-] = 2[SrF_2]$ $= 2 \times 2.71 \times 10^{-4}$ mol/L $= 1.94 \times 10^{-3}$ mol/L
Write the expression for the K_{sp}. Substitute the ion concentrations into the K_{sp} expression and solve.	$K_{sp} = [Sr^{2+}][F^-]^2$ $= (9.71 \times 10^{-4})(1.94 \times 10^{-3})^2$ $= 3.65 \times 10^{-9}$

Check Your Solution
The calculated value of K_{sp} has the correct number of significant digits that agrees with the given data. The small value is expected for a compound of low solubility.

128. Will a precipitate form if 1.00 mL of a 0.100 mol/L silver nitrate solution, $AgNO_3(aq)$, is added to 1.00 L of a 1.00×10^{-5} mol/L solution of sodium chloride, NaCl(aq)? Show your calculations.

What Is Required?
You need to determine whether a precipitate will form.

What Is Given?
You know that 1.00 mL of a 0.100 mol/L silver nitrate solution, $AgNO_3(aq)$, is added to 1.00 L of a 1.00×10^{-5} mol/L solution of sodium chloride, NaCl(aq). The only possible precipitate that can form is silver chloride, AgCl(s).
From Appendix B, you know the K_{sp} for silver chloride is 1.77×10^{-10}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the solubility equilibrium.	$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$
Use the formula $n = cV$ to determine the amount in moles, n, of $Ag^+(aq)$ and $Cl^-(aq)$ in the original solution.	$n_{Ag^+} = cV$ $= 0.100 \text{ mol/}\cancel{L} \times 0.001 \cancel{L}$ $= 1.00 \times 10^{-4} \text{ mol}$ $n_{Cl^-} = cV$ $= 1.00 \times 10^{-5} \text{ mol/}\cancel{L} \times 1.00 \cancel{L}$ $= 1.00 \times 10^{-5} \text{ mol}$
Determine the total volume, V, after mixing the two solutions.	V = 1.00 L + 1.00 mL = 1.00 L + 0.001 L = 1.001 L
Use the formula $c = \frac{n}{V}$ to calculate the molar concentrations, c, of $Ag^+(aq)$ and $Cl^-(aq)$.	$[Ag^+] = \frac{n}{V}$ $= \frac{1.00 \times 10^{-4} \text{ mol}}{1.001 \text{ L}}$ $= 9.99 \times 10^{-5} \text{ mol/L}$ $[Cl^-] = \frac{n}{V}$ $= \frac{1.00 \times 10^{-5} \text{ mol}}{1.001 \text{ L}}$ $= 9.99 \times 10^{-5} \text{ mol/L}$

Write the expression for the K_{sp} for AgCl.	$Q_{sp} = [Ag^+][Cl^-]$
Use the molar concentrations calculated after mixing to determine a trial value Q_{sp}. If $Q_{sp} > K_{sp}$, a precipitate of AgCl(s) forms.	$Q_{sp} = [Ag^+][Cl^-]$ $= 9.99 \times 10^{-5} \times 9.99 \times 10^{-5}$ $= 9.98 \times 10^{-10}$ K_{sp} for AgCl is 1.77×10^{-10}. $Q_{sp} > K_{sp}$ A precipitate will form.

Check Your Solution

The calculations seem reasonable and the correct number of significant digits has been used.