Find the solubility of magnesium fluoride in distilled water at 25C. Report the solubility in units of parts per million (mg/L).

From Appendix J: MgF_2 ; $K_{sp} = 6.4 \times 10^{-9}$

Make an equilibrium chart.

$MgF_2(s) \rightleftharpoons Mg^{2+}(aq) + 2F^-(aq)$; $K_{sp} = [Mg^{2+}][F^-]^2 = 6.4 \times 10^{-9}$

Species	[Initial]	Δ	[Equilibrium]
Mg^{2+}	0	+x	x
F^-	0	+2x	2x

Let "x" equal the change in magnesium ion concentration.

$$[Mg^{2+}][F^-]^2 = 6.4 \times 10^{-9}$$

$$(x)(2x)^2 = 6.4 \times 10^{-9}$$

$$4x^3 = 6.4 \times 10^{-9}$$

$$x = 0.001169607 = [Mg^{2+}]$$

MgF_2: Mg − 1 × 24.31
F − 2 × 19.00
62.31 g/mol

Every time a magnesium fluoride unit dissolves, it produces a magnesium ion. So, the magnesium ion concentration is equal to the concentration of dissolved magnesium fluoride. $[MgF_2]_{dissolved} = [Mg^{2+}] = 0.001169607\ M$

$$\frac{0.001169607 \text{ mol } MgF_2}{L} \times \frac{62.31 \text{ g } MgF_2}{\text{mol } MgF_2} = 0.073 \frac{\text{g } MgF_2}{L} = \boxed{73 \frac{\text{mg } MgF_2}{L}}$$

2 Find the solubility of magnesium fluoride in 0.025 M sodium fluoride solution at 25C. Report the solubility in units of parts per million (mg/L).

$MgF_2(s) \rightleftharpoons Mg^{2+}(aq) + 2F^-(aq)$; $K_{sp} = [Mg^{2+}][F^-]^2 = 6.4 \times 10^{-9}$

Set up an equilibrium chart.

Species	[Initial]	Δ	[Equilibrium]
Mg^{2+}	0	$+x$	x
F^-	0.025	$+2x$	$0.025 + 2x$

Let "x" equal the change in magnesium ion concentration

$(x)(0.025 + 2x)^2 = 6.4 \times 10^{-9}$

$2x \ll 0.025$; $0.025 + 2x \approx 0.025$

$(x)(0.025)^2 = 6.4 \times 10^{-9}$

$x = 1.024 \times 10^{-5} = [Mg^{2+}] = [MgF_2]_{dissolved}$

$$\frac{1.024 \times 10^{-5} \text{ mol } MgF_2}{L} \times \frac{62.31 \text{ g } MgF_2}{\text{mol } MgF_2} = 6.4 \times 10^{-4} \frac{\text{g } MgF_2}{L} = \boxed{0.64 \frac{\text{mg } MgF_2}{L}}$$

3 Will a solution containing 0.00075 M magnesium ion and 0.000069 M fluoride ion precipitate?

Use reaction quotient, Q: $Q = [Mg^{2+}][F^-]^2$

$$Q = (0.00075)(0.000069)^2 = 3.57 \times 10^{-12}$$

Compare Q to Kc:

$$MgF_2(s) \rightleftharpoons Mg^{2+}(aq) + 2F^-(aq)$$

$$Q < Kc$$

$$3.57 \times 10^{-12} < 6.4 \times 10^{-9}$$

Since $Q < Kc$, $\boxed{\text{no precipitation}}$

Q would need to increase in order to reach equilibrium. Since Q depends on the concentration of dissolved ions, that means MORE solid would need to dissolve, so we won't see solid forming here.