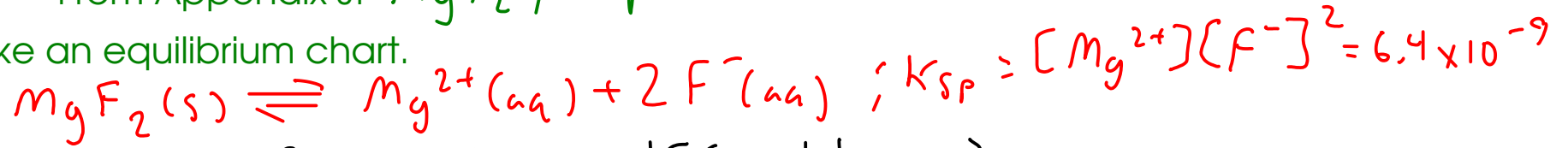


Find the solubility of magnesium fluoride in distilled water at 25C. Report the solubility in units of parts per million (mg/L).

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

Make an equilibrium chart.



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

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

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

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

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

$$x = 0.0011696071 = [Mg^{2+}]$$

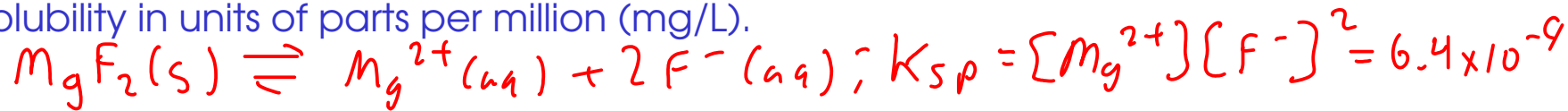
$$\begin{array}{r} MgF_2: Mg - 1 \times 24.31 \\ F - 2 \times 19.00 \\ \hline 62.31 \text{ g/mol} \end{array}$$

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.

$$[MgF_2]_{\text{dissolved}} = [Mg^{2+}] = 0.0011696071 \text{ M}$$

$$\frac{0.0011696071 \text{ mol } MgF_2}{L} \times \frac{62.31 \text{ g } MgF_2}{\text{mol } MgF_2} = 0.073 \frac{\text{g } MgF_2}{L} = \boxed{\frac{73 \text{ mg } MgF_2}{L}}$$

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).



Set up an equilibrium chart.

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

Let "x" equal the change in magnesium ion concentration

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

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

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

$$x = 1.024 \times 10^{-5} = [\text{Mg}^{2+}] = [\text{MgF}_2]_{\text{dissolved}}$$

$$\frac{1.024 \times 10^{-5} \text{ mol MgF}_2}{\text{L}} \times \frac{62.31 \text{ g MgF}_2}{\text{mol MgF}_2} = 6.4 \times 10^{-4} \text{ g MgF}_2 = \frac{0.64 \text{ mg MgF}_2}{\text{L}}$$

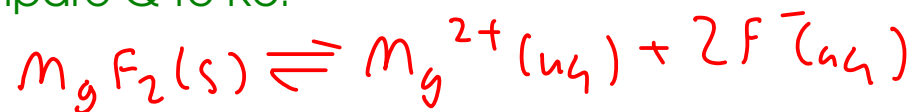
3

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:



$$Q < K_c$$

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

Since  $Q < K_c$ , 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.