Section 7.6 Common Ion Effect

The common ion effect is another example of Le Châtelier's Principle

The solubility of a sparingly soluble salt is reduced in a solution that contains an ion in common with that salt. For instance, the solubility of silver chloride in water is reduced if a solution of sodium chloride is added to a suspension of silver chloride in water

If we have a barium sulfate solution, the solid salt is in equilibrium with its ions:

$$BaSO^{4}{}_{(s)} \rightleftarrows Ba^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$$

If we then add solid barium chloride to this solution, which dissolves to produce Ba^{2+} and Cl^- ions, we are increasing the concentration of Ba^{2+} ions in our solution. (The new Cl^- ions will remain in solution as spectator ions). Ba^{2+} is the ion common to both solutions.

Le Châtelier's Principle tells us that if the concentration of one of the reaction participants is increased, then equilibrium will shift to use up the additional substance.

So adding more Ba^{2+} will force the equilibrium to shift to the left (the reverse direction) in order to use up the added Ba^{2+} ions, producing more solid $BaSO_4$. The concentration of $SO_4{}^{2-}$ will decrease, indicating that solubility has decreased.

A practical example used very widely in areas drawing drinking water from chalk or limestone aquifers. The addition of sodium carbonate to the raw water reduces the hardness of the water. In the water treatment process, highly soluble sodium carbonate salt is added to precipitate out sparingly soluble calcium carbonate. The very pure and finely divided precipitate of calcium carbonate that is generated is a valuable by-product used in the manufacture of toothpaste.

Example: What is the solubility of $PbCl_2$ in 0.10 M NaCl? K_{sp} for $PbCl_2$ is 1.7×10^{-5}.

Solution: Set up the problem as a solubility problem: first, write down the balanced chemical equation and the K_{sp} expression

$PbCl_2(s) <=> Pb^{+2}(aq) + 2Cl^-(aq)$

$K_{sp} = [Pb^{+2}][Cl^-]^2 = 1.7 \times 10^{-5}$

For each mole of that dissolves, 1 mole of and two moles of are formed.

	$PbCl_2$	⇔	Pb^{+2}	$2Cl^-$
Initial			0	0.1
Change			+x	+2x
Equilibrium			x	2x + 0.1

$K_{sp} = [Pb^{+2}][Cl^-]^2$
$1.8 \times 10^{-5} = (x)(2x + 0.1)^2$

If s is small, as we expect, 2x + 0.1 ~ 0.1, so we can try to substitute this in
$1.8 \times 10^{-5} = (x)(0.1)^2$
$x = 1.8 \times 10^{-3}$

This is indeed much smaller than 0.1, so the approximation 2x+0.1 ~ 0.1 is valid. Thus, **the molar solubility of $PbCl_2$ in 0.1M NaCl is 1.8×10^{-3} mol/L**. (For comparison, it is 1.6×10^{-2} mol/L in pure water- it's much less soluble in salt water.)