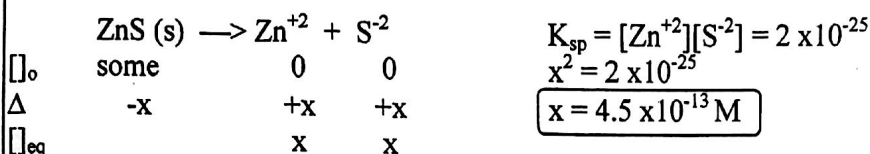


Common Ion Effect

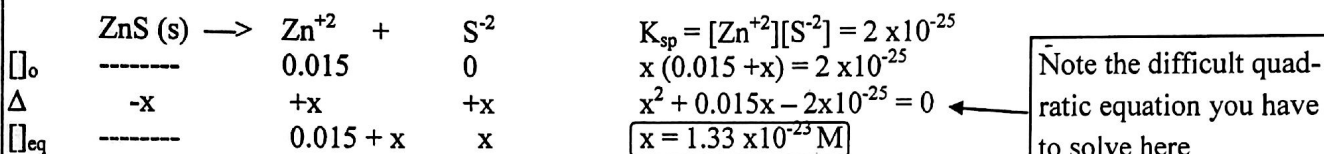
It is more difficult for an ionic compound to dissolve in a solution that already contains one of the ions that is trying to dissolve. Let's prove this

Example 1: Calculate the molar solubility in pure water of ZnS. $K_{sp} = 2 \times 10^{-25}$

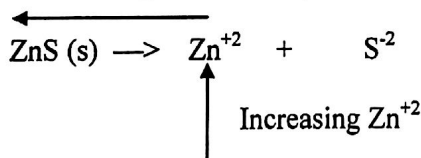


Example 2: Calculate the molar solubility of ZnS in a 0.015 M solution of ZnCl₂.

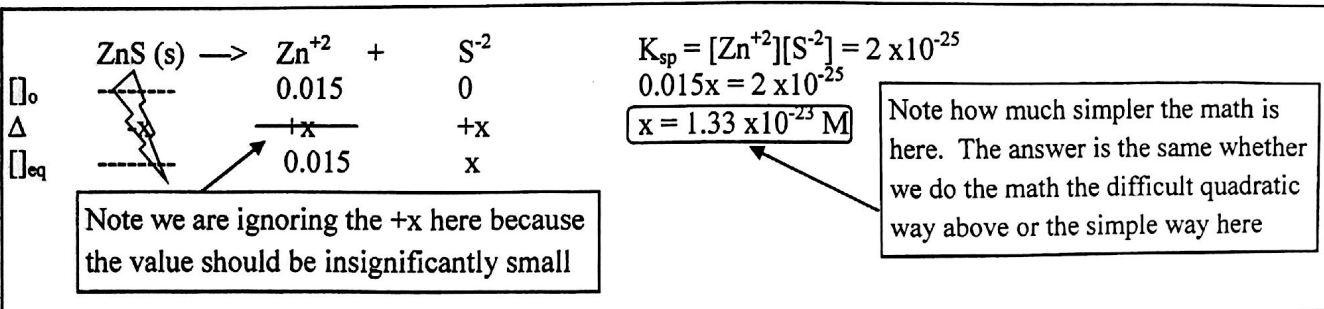
The important thing to note here is that the solid ZnS is trying to dissolve in a solution that already contains one of the ions in the K_{sp} equilibrium. In a 0.015 M solution of ZnCl₂, we are interested in the Zn⁺² ion that is present. Since there is only 1 ion of Zn⁺² for every ZnCl₂, the initial concentration of Zn⁺² is not zero but rather 0.015 M.



There are two things to notice from these two examples. First, the solubility (x) dropped by about 10 billion times from pure water to the solution that already had Zn⁺² ions in it. This should not be surprising though, especially if you consider LeChatelier's Principle. By raising the amount of Zn⁺² ions on the right, it sends the reaction to the left-hand side, thus decreasing the solubility.



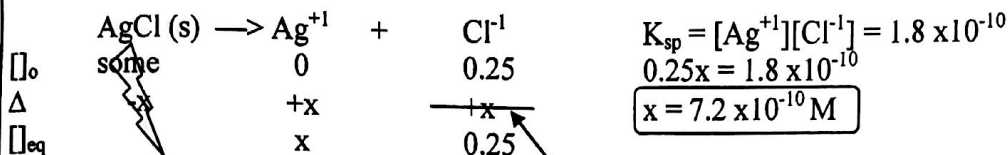
But the second and more mathematically important part for us was that the evil quadratic equation above was not really even necessary. Notice how the answer was 1.33×10^{-23} . This is such a wildly small number compared to the 0.015 M solution the Zn⁺² ion started with it was basically insignificant. If that is the case, we should be able to mathematically ignore it and make our mathematical lives easier. Let's take a look:



Common Ion Effect: If $[x]_o \gg K_{sp}$ then $\Delta[x]_o = 0$ so can be ignored.

As long as the initial value of an ion in a K_{sp} equation is at least 10^3 or 1000 times bigger than K_{sp} then the change for that ion will be insignificant and can be ignored

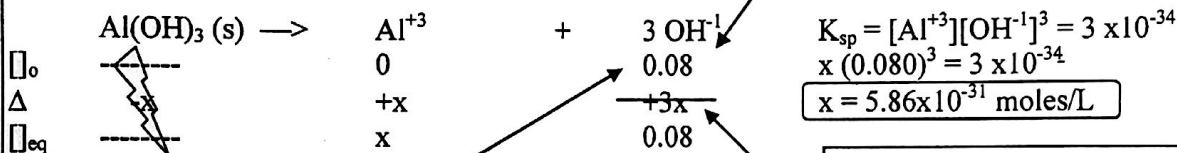
Example 3: Calculate the molar solubility of AgCl in 0.25 M NaCl. $K_{sp} = 1.8 \times 10^{-10}$
 What part of the NaCl are we interested in? The Cl^- part of the NaCl. Since there is only 1 Cl^- for every NaCl, the starting concentration of Cl^- in our K_{sp} equation is 0.25 M



Note how we are ignoring the +x here because the 0.25 M is at least 10^3 times bigger than K_{sp} (1.8×10^{-10})

Example 4: How many g of Al(OH)_3 ($K_{sp} = 3 \times 10^{-34}$) will dissolve in 500 mL of 0.040 M Ba(OH)_2

The first thing to recognize here is that the part of Ba(OH)_2 we are interested in the OH^- ions. Since the formula is Ba(OH)_2 , we must double the 0.040 M to 0.080 M because there are 2 OH^- ions for every Ba(OH)_2 . Thus the initial amount of OH^- ions is 0.080 M.



Note that the 0.08 is the starting amount of the OH^- because of the soluble 0.04 M Ba(OH)_2 solution.

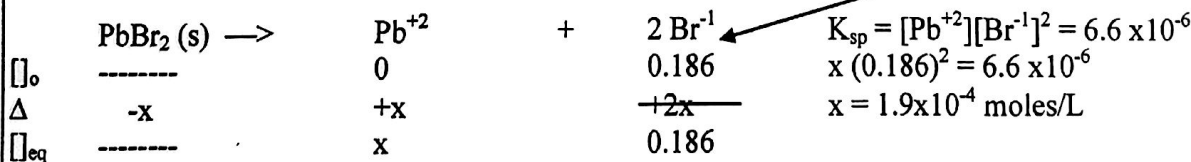
Note that the "3" for the OH^- ion is being taken care of here for the +3x. This is being ignored due to the common ion effect, however.

The solubility of Al(OH)_3 in this problem is 5.86×10^{-31} moles/L. Taking into account the mass of Al(OH)_3 is 78 g/mole and there are 500 mL of solution we get:

$$\frac{5.86 \times 10^{-31} \text{ moles}}{1 \text{ L}} \times \frac{78 \text{ grams}}{\text{mole}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{500 \text{ mL}}{1} = 2.29 \times 10^{-29} \text{ g Al(OH)}_3$$

Example 5: How many mg of PbBr_2 ($K_{sp} = 6.6 \times 10^{-6}$) will dissolve in 100 mL of 0.062 M AlBr_3

We are interested in the Br^- part of AlBr_3 so the $[\text{Br}^-] = 3 * 0.062 \text{ M} = 0.186 \text{ M Br}^-$ initially



The solubility of PbBr_2 in this problem is 1.9×10^{-4} moles/L. Taking into account the mass of PbBr_2 is 367 g/mole and there are 100 mL of solution we get:

$$\frac{1.9 \times 10^{-4} \text{ moles}}{1 \text{ L}} \times \frac{367 \text{ grams}}{\text{mole}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{100 \text{ mL}}{1} \times \frac{1000 \text{ mg}}{\text{g}} = 6.97 \text{ mg PbBr}_2$$