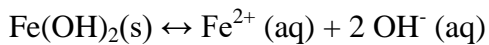
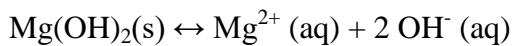


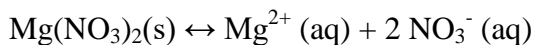
Piece 1 Write a balanced equation for the dissolution of  $\text{Fe}(\text{OH})_2$  (s) in water.



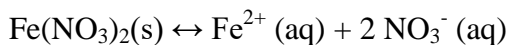
Piece 2 Write a balanced equation for the dissolution of  $\text{Mg}(\text{OH})_2$  (s) in water.



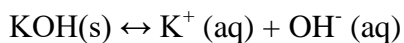
Piece 3: Write a balanced equation for the dissolution of  $\text{Mg}(\text{NO}_3)_2$  (s) in water.



Piece 4: Write a balanced equation for the dissolution of  $\text{Fe}(\text{NO}_3)_2$  (s) in water.



Piece 5: Write a balanced equation for the dissolution of  $\text{KOH}$  (s) in water.



Piece 6: If I dissolve  $\text{Mg}(\text{NO}_3)_2$ ,  $\text{Fe}(\text{NO}_3)_2$ , and  $\text{KOH}$  in water, what species are present in my solution?



Piece 7: Given the following  $K_{sp}$  values, which salt is LEAST soluble in water?

$$K_{sp}[\text{Fe}(\text{OH})_2] = 4.9 \times 10^{-17}$$

$$K_{sp}[\text{Mg}(\text{OH})_2] = 5.6 \times 10^{-12}$$

$$K_{sp}[\text{Mg}(\text{NO}_3)_2] = 3.14 \times 10^4$$

$$K_{sp}[\text{Fe}(\text{NO}_3)_2] = 6.28 \times 10^3$$

$$K_{sp}[\text{KOH}] = 1.26 \times 10^5$$

$$K_{sp}[\text{Fe}(\text{OH})_2] = 4.9 \times 10^{-17}$$

Smaller  $K_{sp}$ , less soluble!

Piece 8. Consider 1.0 L of solution that is  $2.3 \times 10^{-2}$  M in  $\text{Fe}(\text{NO}_3)_2$  and  $1.5 \times 10^{-2}$  M in  $\text{Mg}(\text{NO}_3)_2$ .

If potassium hydroxide is used to selectively precipitate the cations, which cation will precipitate first?

$\text{Fe}^{2+}$  because  $\text{Fe}(\text{OH})_2$  is least soluble!

Puzzle 1. What minimum concentration of KOH is required to cause the precipitation of the cation that precipitates first? How much of the first cation is still in solution when the second cation begins to precipitate?

	Fe(OH) <sub>2</sub> (s)	↔	Fe <sup>2+</sup> (aq)	+ 2 OH <sup>-</sup> (aq)
I			0.023	Add this
C	+x		-x	-2x
E			0.023-x	Added-2x

$$K_{sp} = [Fe^{2+}][OH^{-}]^2 = 4.9 \times 10^{-17}$$

Nothing happens until the OH<sup>-</sup> concentration gets high enough for the Q<sub>sp</sub> to equal the K<sub>sp</sub>. So, I can just calculate when that happens by keeping the Fe<sup>2+</sup> the same and adding OH<sup>-</sup>:

$$[0.023 M][OH^{-}]^2 = 4.9 \times 10^{-17}$$

$$[OH^{-}] = 4.62 \times 10^{-8} M$$

So, if I keep adding OH<sup>-</sup>, the Fe<sup>2+</sup> starts precipitating as Fe(OH)<sub>2</sub> to keep the K<sub>sp</sub> satisfied. So, over time, the [OH<sup>-</sup>] creeps up while the [Fe<sup>2+</sup>] creeps down. Eventually, if the [OH<sup>-</sup>] gets too big, the Mg<sup>2+</sup> will start precipitating.

When? When the K<sub>sp</sub> of Mg(OH)<sub>2</sub> gets violated.

$$K_{sp} = [Mg^{2+}][OH^{-}]^2 = 5.6 \times 10^{-12}$$

$$[0.015 M][OH^{-}]^2 = 5.6 \times 10^{-12}$$

$$[OH^{-}] = 1.93 \times 10^{-5} M$$

The [OH<sup>-</sup>] can only get this high if the [Fe<sup>2+</sup>] has dropped enough to allow it to stay in solution.

$$K_{sp} = [Fe^{2+}][OH^{-}]^2 = 4.9 \times 10^{-17}$$

$$[Fe^{2+}][1.93 \times 10^{-5} M]^2 = 4.9 \times 10^{-17}$$

$$[Fe^{2+}] = 1.32 \times 10^{-7} M$$

Puzzle 2. What is the total mass of KOH that must be added to get the second cation to begin to precipitate?

When the  $Mg^{2+}$  starts precipitating, there is some  $OH^-$  in solution AND there is some  $OH^-$  that is now bound into the precipitate,  $Fe(OH)_2$ . So, the total amount is the amount of  $OH^-$  in solution and the amount of  $OH^-$  in the precipitate.

We already calculated the amount in solution in Puzzle 1:

$[OH^-]=1.93 \times 10^{-5} M$  when the  $Mg^{2+}$  starts to precipitate.

$$1.93 \times 10^{-5} M = \frac{1.93 \times 10^{-5} \text{ mol } OH^-}{L} \frac{1 \text{ mol } KOH}{1 \text{ mol } OH^-} = \frac{1.93 \times 10^{-5} \text{ mol } KOH}{L}$$

$$\frac{1.93 \times 10^{-5} \text{ mol } KOH}{L} 1.00 L = 1.9310^{-5} \text{ mol } KOH \frac{56.1 \text{ g } KOH}{\text{mol } KOH}$$

$$= 1.083 \times 10^{-3} \text{ g } KOH \text{ in solution}$$

For the precipitate, we need to look at how much  $Fe^{2+}$  precipitated:

$0.023 M$  initial -  $1.32 \times 10^{-7} M = 0.0229999 M$  precipitate (almost all of it!)

$$1.0 L \frac{0.023 \text{ mol } Fe^{2+}}{L} \frac{2 \text{ mol } OH^-}{1 \text{ mol } Fe^{2+}} \frac{1 \text{ mol } KOH}{1 \text{ mol } OH^-} \frac{56.1 \text{ g } KOH}{\text{mol } KOH} = 2.581 \text{ g } KOH \text{ in ppt}$$

Total amount of KOH added =  $0.00108 \text{ g}$  in solution +  $2.581 \text{ g}$  in ppt =  $2.582 \text{ g}$  KOH total