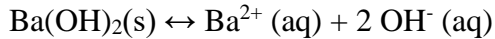
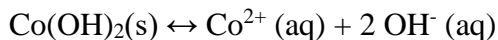


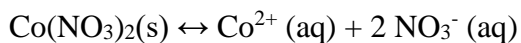
Piece 1 Write a balanced equation for the dissolution of $\text{Ba}(\text{OH})_2$ (s) in water.



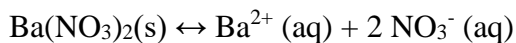
Piece 2 Write a balanced equation for the dissolution of $\text{Co}(\text{OH})_2$ (s) in water.



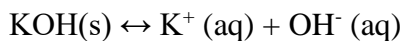
Piece 3: Write a balanced equation for the dissolution of $\text{Co}(\text{NO}_3)_2$ (s) in water.



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



Piece 5: Write a balanced equation for the dissolution of KOH (s) in water.



Piece 6: If I dissolve $\text{Co}(\text{NO}_3)_2$, $\text{Ba}(\text{NO}_3)_2$, and 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?

$$K_{sp}[\text{Ba}(\text{OH})_2] = 5.0 \times 10^{-3}$$

$$K_{sp}[\text{Co}(\text{OH})_2] = 5.9 \times 10^{-15}$$

$$K_{sp}[\text{Ba}(\text{NO}_3)_2] = 3.14 \times 10^2$$

$$K_{sp}[\text{Co}(\text{NO}_3)_2] = 6.28 \times 10^1$$

$$K_{sp}[\text{KOH}] = 1.26 \times 10^2$$

$$K_{sp}[\text{Co}(\text{OH})_2] = 5.9 \times 10^{-15}$$

Smaller K_{sp} , less soluble!

Piece 8. Consider 1.0 L of solution that is 0.100 M in $\text{Ba}(\text{NO}_3)_2$ and 0.100 M in $\text{Co}(\text{NO}_3)_2$.

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

Co^{2+} because $\text{Co}(\text{OH})_2$ 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?

	Co(OH) ₂ (s)	↔	Co ²⁺ (aq)	+ 2 OH ⁻ (aq)
I			0.100	Add this
C	+x		-x	-2x
E			0.100-x	Added-2x

$$K_{sp} = [Co^{2+}][OH^{-}]^2 = 5.9 \times 10^{-15}$$

Nothing happens until the OH⁻ concentration gets high enough for the Q_{sp} to equal the K_{sp}. So, I can just calculate when that happens by keeping the Co²⁺ the same and adding OH⁻:

$$[0.100 M][OH^{-}]^2 = 5.9 \times 10^{-15}$$

$$[OH^{-}] = 2.42 \times 10^{-7} M$$

So, if I keep adding OH⁻, the Co²⁺ starts precipitating as Co(OH)₂ to keep the K_{sp} satisfied. So, over time, the [OH⁻] creeps up while the [Co²⁺] creeps down. Eventually, if the [OH⁻] gets too big, the Ba²⁺ will start precipitating.

When? When the K_{sp} of Ba(OH)₂ gets violated.

$$K_{sp} = [Ba^{2+}][OH^{-}]^2 = 5.0 \times 10^{-3}$$

$$[0.100 M][OH^{-}]^2 = 5.0 \times 10^{-3}$$

$$[OH^{-}] = 0.2236 M$$

The [OH⁻] can only get this high if the [Co²⁺] has dropped enough to allow it to stay in solution.

$$K_{sp} = [Co^{2+}][OH^{-}]^2 = 5.9 \times 10^{-15}$$

$$[Co^{2+}][0.2236 M]^2 = 5.9 \times 10^{-15}$$

$$[Co^{2+}] = 1.18 \times 10^{-13} M$$

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

When the Ba^{2+} starts precipitating, there is some OH^- in solution AND there is some OH^- that is now bound into the precipitate, $\text{Co}(\text{OH})_2$. 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:

$[\text{OH}^-]=0.2236 \text{ M}$ when the Ba^{2+} starts to precipitate.

$$1 \text{ L} \frac{0.2236 \text{ mol OH}^-}{\text{L}} \frac{1 \text{ mol KOH}}{1 \text{ mol OH}^-} = 0.2236 \text{ mol KOH} \frac{56.1 \text{ g KOH}}{\text{mol KOH}} \\ = 12.54 \text{ g KOH in solution}$$

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

$0.100 \text{ M initial} - 1.18 \times 10^{-13} \text{ M} = 0.100 \text{ M precipitate}$ (almost all of it!)

$$1.0 \text{ L} \frac{0.100 \text{ mol Co}^{2+}}{\text{L}} \frac{2 \text{ mol OH}^-}{1 \text{ mol Co}^{2+}} \frac{1 \text{ mol KOH}}{1 \text{ mol OH}^-} \frac{56.1 \text{ g KOH}}{\text{mol KOH}} = 11.22 \text{ g KOH in ppt}$$

Total amount of KOH added = $12.54 \text{ g in solution} + 11.22 \text{ g in ppt} = 23.76 \text{ g KOH total}$