

I'm human





Predicting the solubility of common inorganic compounds and writing net ionic equations for precipitation reactions are key learning objectives. Precipitation reactions involve the formation of one or more solid products from dissolved substances, often through ion exchange between ionic compounds in aqueous solution. These reactions occur naturally and have significant industrial applications, as well as playing a crucial role in chemical analysis techniques. The major species involved in a reaction between magnesium hydroxide and ammonium ions are ammonia and water. Adding these to the chemical equation shows how they interact with each other and their surroundings. The formation of solid lead iodide occurs when potassium iodide and lead nitrate solutions are mixed, producing a bright yellow precipitate known as iodine yellow. This reaction is consistent with the solubility guidelines, which state that only insoluble combinations between lead and iodide can occur. Net ionic equations, such as  $\text{Pb}^{2+}(\text{aq}) + 2\text{I}^{-}(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$ , describe these chemical interactions in a more detailed manner. The characteristics of pure crystals make them useful for the fabrication of X-ray and gamma ray detectors. The solubility guidelines listed in Table [1](#) can be used to predict whether a precipitation reaction will occur when solutions containing soluble ionic compounds are mixed together. By identifying all the ions present in the solution and considering their possible cation/anion pairing, one can determine if an insoluble compound could form. For example, combining silver nitrate and sodium chloride solutions produces a mixture containing  $\text{Ag}^{+}$ ,  $\text{NO}_3^{-}$ ,  $\text{Na}^{+}$ , and  $\text{Cl}^{-}$  ions. Based on the solubility guidelines, only  $\text{AgCl}$  is an insoluble combination among these ions. Therefore, a precipitation reaction occurs when the following net ionic equation is considered:  $\text{Ag}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq}) \rightarrow \text{AgCl}(\text{s})$ . In another scenario, mixing solutions of potassium sulfate and barium nitrate will yield a solution containing  $\text{K}^{+}$ ,  $\text{NO}_3^{-}$ ,  $\text{Ba}^{2+}$ , and  $\text{SO}_4^{2-}$  ions. Considering possible cation/anion pairing based on the solubility guidelines indicates that  $\text{BaSO}_4$  is insoluble. As a result, a precipitation reaction occurs when the following net ionic equation is examined:  $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$ . In addition, combining lithium chloride and silver acetate solutions results in a mixture containing  $\text{Li}^{+}$ ,  $\text{C}_2\text{H}_3\text{O}_2^{-}$ ,  $\text{Ag}^{+}$ , and  $\text{Cl}^{-}$  ions. Based on the solubility guidelines, only  $\text{AgCl}$  is an insoluble combination among these ions. As such, a precipitation reaction occurs when the following net ionic equation is considered:  $\text{Ag}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq}) \rightarrow \text{AgCl}(\text{s})$ . Lastly, mixing lead nitrate and ammonium carbonate solutions produces a mixture containing  $\text{Pb}^{2+}$ ,  $\text{NO}_3^{-}$ ,  $\text{NH}_4^{+}$ , and  $\text{CO}_3^{2-}$  ions. Considering possible cation/anion pairing based on the solubility guidelines indicates that  $\text{PbCO}_3$  is insoluble. As a result, a precipitation reaction occurs when the following net ionic equation is examined:  $\text{Pb}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{PbCO}_3(\text{s})$ . Note that these examples illustrate how to predict and write balanced net ionic equations for precipitation reactions using the solubility guidelines. The suitable reagent from sodium chloride, sodium hydroxide, or sodium sulfate for precipitating barium ions in water is sodium sulfate. The resulting compound's formula would be  $\text{BaSO}_4$ . Although lab experiments involving aqueous silver nitrate and sodium sulfate didn't yield a visible precipitate, the chemical equation supports the formation of  $\text{Ag}_2\text{SO}_4$  (silver sulfate). A possible explanation for this discrepancy lies in the fact that silver sulfate may have been decomposed by light, similar to some photographic film substances. Wikipedia mentions an association with "ruby red illumination," but further research is needed to confirm this theory. The solubility classification of silver sulfate as "sparingly soluble" also remains unclear without additional context.

Does silver nitrate form a precipitate. Does silver sulfate precipitate. Does a precipitate form when silver nitrate and sodium sulfate are mixed. Silver nitrate to silver. Does silver nitrate and sodium chloride form a precipitate.