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If 15.5 ML Of 0.225 M Aqueous Magnesium Chloride Is Added To 37.5 ML Of 0.250 M Aqueous Lead(II) Nitrate, the resulting chemical interactions primarily involve potential double displacement reactions, leading to the formation of insoluble precipitates and dissolved ions in solution. This scenario is a classic example of a precipitation reaction in aqueous solutions, which is fundamental in understanding solubility rules, reaction stoichiometry, and net ionic equations. In this article, we will explore the step-by-step process of analyzing such a reaction, including calculating initial moles, predicting precipitate formation, and understanding the implications of solubility equilibria.

Understanding the Reactants and Their Properties

Magnesium Chloride ($MgCl_2$)

- Magnesium chloride is an ionic compound composed of Mg^{2+} cations and Cl^- anions.
- It is highly soluble in water, dissociating completely into Mg^{2+} and Cl^- ions.
- Molar mass of $MgCl_2$: approximately 95.21 g/mol.

Lead(II) Nitrate ($Pb(NO_3)_2$)

- Lead(II) nitrate dissociates into Pb^{2+} cations and NO_3^- anions.
- It is also highly soluble in water.
- Molar mass of $Pb(NO_3)_2$: approximately 331.2 g/mol.

Calculating the Initial Moles of Reactants

Before predicting the reaction, we need to determine the number of moles of each reactant present in the solutions.

Moles of Magnesium Chloride

- Volume of $MgCl_2$ solution: 15.5 mL = 0.0155 L.
- Molarity of $MgCl_2$: 0.225 mol/L.

Calculation:

$$
\text{Moles of MgCl}_2 = 0.225\, \text{mol/L} \times 0.0155\, \text{L} = 0.0034875\, \text{mol}
$$

Moles of Lead(II) Nitrate

- Volume of $Pb(NO_3)_2$ solution: 37.5 mL = 0.0375 L.
- Molarity of $Pb(NO_3)_2$: 0.250 mol/L.

Calculation:

$$
\text{Moles of Pb(NO}_3)_2 = 0.250\, \text{mol/L} \times 0.0375\, \text{L} = 0.009375\, \text{mol}
$$

Predicting the Reaction: Double Displacement and Precipitation

Possible Double Displacement Reaction

- When solutions of magnesium chloride and lead(II) nitrate are mixed, a double displacement reaction can occur:

$$
\text{MgCl}_2 (aq) + \text{Pb(NO}_3)_2 (aq) \rightarrow \text{PbCl}_2 (s) + \text{Mg(NO}_3)_2 (aq)
$$

- The reaction involves the exchange of anions between the cations.

Solubility Rules and Precipitate Formation

- To determine if a precipitate forms, consult solubility rules:
- Most chloride salts are soluble, except those of Ag^+, Pb^{2+}, and Hg_2^{2+}.
- Lead(II) chloride ($PbCl_2$) is insoluble in water.
- Magnesium nitrate ($Mg(NO_3)_2$) is soluble.

- Therefore, $PbCl_2$ will precipitate out if the product of the ionic concentrations exceeds its solubility product (Ksp).

Calculating the Limiting Reactant and Extent of Reaction

Reaction Stoichiometry

- The balanced equation indicates a 1:1 molar ratio between $MgCl_2$ and $Pb(NO_3)_2$ for precipitate formation.

Comparison of Moles

- Moles of $MgCl_2$: 0.0034875 mol
- Moles of $Pb(NO_3)_2$: 0.009375 mol

- Since $MgCl_2$ is in lesser molar amount, it is the limiting reactant.

Extent of Precipitate Formation

- All $MgCl_2$ will react if enough $Pb(NO_3)_2$ is present.
- The maximum moles of $PbCl_2$ precipitated:

$$
\text{Moles of PbCl}_2 = \text{Moles of MgCl}_2 = 0.0034875\, \text{mol}
$$

- Remaining moles of $Pb(NO_3)_2$ after reaction:

$$
0.009375\, \text{mol} - 0.0034875\, \text{mol} = 0.0058875\, \text{mol}
$$

Calculating Concentrations After Reaction

Volume of the Mixture

- Total volume:

$$
15.5\, \text{mL} + 37.5\, \text{mL} = 53\, \text{mL} = 0.053\, \text{L}
$$

Concentrations of Ions in Solution

- Remaining Mg^{2+}:

$$
\frac{0.0034875\, \text{mol}}{0.053\, \text{L}} \approx 0.0658\, \text{M}
$$

- Remaining NO_3^- (from unreacted $Pb(NO_3)_2$):

$$
\frac{0.0058875\, \text{mol}}{0.053\, \text{L}} \approx 0.1111\, \text{M}
$$

- Pb^{2+} ions after reaction:

$$
\text{Initial} - \text{reacted} = 0.009375\, \text{mol} - 0.0034875\, \text{mol} = 0.0058875\, \text{mol}
$$

Concentration:

$$
\frac{0.0058875\, \text{mol}}{0.053\, \text{L}} \approx 0.1111\, \text{M}
$$

- Cl^- ions from $MgCl_2$:

$$
\text{Moles of Cl}^- = 2 \times 0.0034875\, \text{mol} = 0.006975\, \text{mol}
$$

Concentration:

$$
\frac{0.006975\, \text{mol}}{0.053\, \text{L}} \approx 0.1317\, \text{M}
$$

Understanding the Precipitation Process and Ksp

Solubility Product (Ksp) of $PbCl_2$

- The Ksp of $PbCl_2$ at 25°C is approximately 1.7×10^{-5}.

Calculating the Ionic Product

- The ionic product for $PbCl_2$:

$$
[\text{Pb}^{2+}] \times [\text{Cl}^-]^2 = 0.1111\, \text{M} \times (0.1317\, \text{M})^2
$$

Calculations:

$$
0.1111 \times (0.1317)^2 \approx 0.1111 \times 0.01733 \approx 0.00193
$$

- Since 0.00193 > 1.7×10^{-5}, the ionic product exceeds the Ksp, indicating that $PbCl_2$ will precipitate until equilibrium is established.

Conclusion: Formation of Precipitate

- The reaction results in the formation of solid $PbCl_2$, which precipitates out of the solution.
- The amount of $PbCl_2$ precipitated corresponds to the limiting reactant $MgCl_2$, i.e., 0.0034875 mol.
- The excess $Pb(NO_3)_2$ remains in solution, providing nitrate ions and unreacted lead ions.

Implications and Applications

Practical Significance

- This type of reaction illustrates the principle of selective precipitation, which is used in qualitative analysis to separate ions.
- It also demonstrates the importance of solubility rules in predicting reaction outcomes in aqueous solutions.
- Understanding precipitation reactions is essential in industries such as water treatment, mineral processing, and chemical synthesis.

Further Considerations

- Temperature dependence of solubility: Ksp values can vary with temperature.
- Effect of ionic strength: High ionic strength can influence solubility.
- Complex formation: In some cases, ions can form soluble complexes, affecting precipitation.

Summary

- The initial moles of reactants are calculated based on given concentrations and volumes.
- $MgCl_2$ is the limiting reactant, determining the amount of $PbCl_2$ precipitated.
- The reaction leads to the formation of insoluble $PbCl_2$, which can be predicted using solub

Frequently Asked Questions

What is the overall reaction when magnesium chloride reacts with lead(II) nitrate in aqueous solution?

The reaction is a double displacement reaction: $MgCl_2 + Pb(NO_3)_2 \rightarrow Mg(NO_3)_2 + PbCl_2(s)$. Lead(II) chloride precipitates out as a solid.

How do you calculate the moles of magnesium chloride added to the solution?

Multiply the volume in liters (15.5 mL = 0.0155 L) by the molarity (0.225 M): 0.0155 L × 0.225 mol/L = 0.0034875 mol of $MgCl_2$.

What is the initial number of moles of lead(II) nitrate in the solution?

Convert volume to liters (37.5 mL = 0.0375 L) and multiply by molarity: 0.0375 L × 0.250 mol/L = 0.009375 mol of $Pb(NO_3)_2$.

Will a precipitate of lead(II) chloride form, and if so, how much can theoretically precipitate?

Yes, since lead(II) chloride is insoluble. The limiting reagent is magnesium chloride, which provides 0.0034875 mol; since the reaction consumes 1 mol of $MgCl_2$ per mol of $PbCl_2$, up to 0.0034875 mol of $PbCl_2$ can precipitate.

What is the final concentration of lead(II) nitrate after the reaction?

Assuming complete reaction and removal of precipitate, the remaining lead(II) nitrate is the initial amount minus the amount reacted; if all of $MgCl_2$ reacts, 0.009375 mol – 0.0034875 mol = 0.0058875 mol of $Pb(NO_3)_2$ remains in the original volume plus any additional volume if mixing is considered. For simplicity, initial concentration remains approximately 0.250 M since the added $MgCl_2$ volume is small.

How do you determine the net ionic equation for this reaction?

The net ionic equation is Pb^{2+}(aq) + 2Cl^-(aq) → $PbCl_2$(s), showing the formation of insoluble lead(II) chloride from lead(II) and chloride ions in solution.

Additional Resources

When aqueous solutions of magnesium chloride and lead(II) nitrate are combined, a fascinating chemical reaction ensues, leading to the formation of an insoluble precipitate and a redistribution of ions within the solution. This process exemplifies key principles of solubility, double displacement reactions, and stoichiometry in aqueous chemistry, offering insight into how ionic interactions govern the outcomes of mixing different compounds. Analyzing the specifics of this reaction provides a deeper understanding of precipitation reactions, their practical applications, and the underlying chemical principles involved.

Introduction and Contextual Background

In the realm of aqueous chemistry, reactions involving the exchange of ions between soluble salts are fundamental to understanding how substances behave when dissolved in water. When solutions of magnesium chloride ($MgCl_2$) and lead(II) nitrate ($Pb(NO_3)_2$) are combined, the ions present in the mixture can interact to produce new compounds, some of which may be insoluble and precipitate out of the solution. Such reactions are classic demonstrations of solubility rules and are of practical importance in fields such as chemical analysis, wastewater treatment, and the synthesis of specific compounds.

This specific scenario involves adding 15.5 mL of a 0.225 M aqueous magnesium chloride solution to 37.5 mL of a 0.250 M aqueous lead(II) nitrate solution. The focus is to analyze what chemical changes occur, determine the precipitate formed, and understand the underlying principles governing this process.

Chemical Reactions and Stoichiometry

Identifying the Relevant Chemical Equation

The key to understanding the reaction lies in the double displacement (metathesis) reaction:

\[\mathrm{MgCl}_2 (aq) + \mathrm{Pb(NO}_3)_2 (aq) \rightarrow \text{Products} \]

Given the ions present, the potential insoluble compound (precipitate) would be lead(II) chloride ($PbCl_2$), as per solubility rules, which state that most chlorides are soluble except those of lead(II), silver, and mercury(I). Lead(II) chloride is known to be sparingly soluble, making it a prime candidate to precipitate out.

The full balanced chemical equation:

$$ \mathrm{MgCl}_2 (aq) + \mathrm{Pb(NO}_3)_2 (aq) \rightarrow \mathrm{PbCl}_2 (s) + \mathrm{Mg(NO}_3)_2 (aq) $$

This indicates that when magnesium chloride reacts with lead(II) nitrate, lead(II) chloride precipitates, and magnesium nitrate remains in solution.

Calculating the Moles of Reactants

To analyze this reaction quantitatively, determine the number of moles of each reactant:

- Magnesium chloride:

$$
\text{Volume} = 15.5\,\text{mL} = 0.0155\,\text{L}
$$

$$
\text{Concentration} = 0.225\,\text{mol/L}
$$

$$
\text{Moles of MgCl}_2 = 0.0155\,\text{L} \times 0.225\,\text{mol/L} = 0.0034875\,\text{mol}
$$

- Lead(II) nitrate:

$$
\text{Volume} = 37.5\,\text{mL} = 0.0375\,\text{L}
$$

$$
\text{Concentration} = 0.250\,\text{mol/L}
$$

$$
\text{Moles of Pb(NO}_3)_2 = 0.0375\,\text{L} \times 0.250\,\text{mol/L} = 0.009375\,\text{mol}
$$

Determining the Limiting Reagent and Precipitate Formation

Stoichiometric Ratios and Reaction Extent

The balanced equation indicates a 1:1 molar ratio between $MgCl_2$ and $Pb(NO_3)_2$, meaning:

- 1 mole of $MgCl_2$ reacts with 1 mole of $Pb(NO_3)_2$ to produce 1 mole of $PbCl_2$ precipitate.

Comparing the moles:

- $MgCl_2$: 0.0034875 mol
- $Pb(NO_3)_2$: 0.009375 mol

Since $MgCl_2$ is present in a smaller molar amount relative to the stoichiometric ratio, magnesium chloride is the limiting reagent.

Implication: The amount of lead(II) chloride precipitated is limited by the amount of magnesium chloride available.

Calculating the Mass and Volume of Precipitate Formed

- Moles of $PbCl_2$ formed: Equivalent to moles of $MgCl_2$, i.e., 0.0034875 mol.

- Molar mass of $PbCl_2$:

$$
\text{Pb} = 207.2\,\text{g/mol}
$$
$$
\text{Cl} = 35.45\,\text{g/mol}
$$
$$
\text{PbCl}_2 = 207.2 + 2 \times 35.45 = 278.1\,\text{g/mol}
$$

- Mass of $PbCl_2$ precipitate:

$$
0.0034875\,\text{mol} \times 278.1\,\text{g/mol} \approx 0.970\,\text{g}
$$

This indicates that approximately 0.97 grams of lead(II) chloride will precipitate out upon mixing.

Solution Concentrations Post-Reaction

Remaining Ions in Solution

Not all lead(II) nitrate will react because it exceeds the amount needed to fully react with the magnesium chloride. After the reaction:

- Moles of $Pb(NO_3)_2$ initially: 0.009375 mol
- Moles reacted: 0.0034875 mol (equivalent to $MgCl_2$)

Remaining:

$$
0.009375\,\text{mol} - 0.0034875\,\text{mol} = 0.0058875\,\text{mol}
$$

Thus, unreacted lead(II) nitrate remains in solution.

- Remaining $MgCl_2$: Fully reacted (limiting reagent).

- Magnesium nitrate in solution:

$$
\text{Moles} = 0.0034875\,\text{mol}
$$

- Mass of magnesium nitrate:

$$
\text{Mg(NO}_3)_2 = 24.3 + 2 \times 62.0 = 148.3\,\text{g/mol}
$$

- Mass:

$$
0.0034875\,\text{mol} \times 148.3\,\text{g/mol} \approx 0.517\,\text{g}
$$

- Remaining lead(II) nitrate:

$$
0.0058875\,\text{mol} \times 331.2\,\text{g/mol} \approx 1.949\,\text{g}
$$

- Concentrations in the final mixture:

The total volume after mixing:

$$
15.5\,\text{mL} + 37.5\,\text{mL} = 53\,\text{mL} = 0.053\,\text{L}
$$

Therefore:

- Concentration of unreacted $Pb(NO_3)_2$:

$$
\frac{0.0058875\,\text{mol}}{0.053\,\text{L}} \approx 0.111\,\text{M}
$$

- Concentration of $Mg(NO_3)_2$:

$$
\frac{0.0034875\,\text{mol}}{0.053\,\text{L}} \approx 0.0658\,\text{M}
$$

Implications and Broader Significance

Precipitation and Its Practical Applications

The formation of $PbCl_2$ highlights the importance of understanding solubility rules in practical contexts. Lead(II) chloride's sparing solubility makes it useful for:

- Purification processes: Precipitating lead compounds for removal from waste streams.
- Analytical chemistry: Qualitative analysis of halide ions.
- Industrial synthesis: Producing specific lead halide compounds.

However, the toxicity of lead compounds necessitates careful handling and disposal, underscoring the importance of safety in laboratory and industrial settings.

Theoretical and Educational Significance

This reaction serves as an educational example demonstrating:

- The concept of limiting reagents and stoichiometry.
- How solubility rules predict precipitate formation.
- The importance of molar calculations in chemical reactions.
- The practical importance of balancing and analyzing chemical equations.

Conclusion and Final Remarks