Chapter 9

Reactions in Aqueous Solutions

Section 9.1 Net Ionic Equations and Qualitative Analysis

Solutions for Practice Problems

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1. Practice Problem (page 410)

Write the net ionic equation for this reaction:

$Ba(ClO_3)_2(aq) + Na_3PO_4(aq) \rightarrow Ba_3(PO_4)_2(s) + NaClO_3(aq)$

What Is Required?

You need to write the net ionic equation for the reaction.

What Is Given?

You know that the reaction between barium chlorate and sodium phosphate is a double displacement reaction.

You know that barium phosphate, $Ba_3(PO_4)_2(s)$, is the precipitate.

You know the skeleton equation:

$Ba(ClO_3)_2(aq) + Na_3PO_4(aq) \rightarrow Ba_3(PO_4)_2(s) + NaClO_3(aq)$

Plan Your Strategy

Write the complete chemical equation for the reaction.

Write $Ba(ClO_3)_2(aq)$, $Na_3PO_4(aq)$, and $NaClO_3(aq)$ as ions. Leave $Ba_3(PO_4)_2(s)$ as a formula unit since this ionic compound has low solubility.

Write the complete ionic equation for the reaction.

Identify the spectator ions, and cancel them on both sides of the equation.

Write the net ionic equation.

Act on Your Strategy

The precipitate is barium phosphate, $Ba_3(PO_4)_2(s)$.

Skeleton equation:

$Ba(ClO_3)_2(aq) + Na_3PO_4(aq) \rightarrow Ba_3(PO_4)_2(s) + NaClO_3(aq)$

Complete chemical equation:

$3Ba(ClO_3)_2(aq) + 2Na_3PO_4(aq) \rightarrow Ba_3(PO_4)_2(s) + 6NaClO_3(aq)$

Complete ionic equation:

$$3Ba^{2+}(aq) + \cancel{6ClO_3^{-}(aq)} + \cancel{6Na^{+}(aq)} + 2PO_4^{3-}(aq) \rightarrow Ba_3(PO_4)_2(s) + \cancel{6\,Na^{+}(aq)} + \cancel{6ClO_3^{-}(aq)}$$

Net ionic equation:
$3Ba^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ba_3(PO_4)_2(s)$

Check Your Solution
The net ionic equation is balanced, including the charges on the ions.

2. Practice Problem (page 410)
Write the net ionic equation for this reaction:

$$Na_2SO_4(aq) + Sr(OH)_2(aq) \rightarrow SrSO_4(s) + NaOH(aq)$$

What Is Required?
You need to write the net ionic equation for the reaction.

What Is Given?
You know that the reaction between sodium sulfate and strontium hydroxide is a double displacement reaction.
You know that strontium sulfate, $SrSO_4(s)$, is the precipitate.
You know the skeleton equation:
$Na_2SO_4(aq) + Sr(OH)_2(aq) \rightarrow SrSO_4(s) + NaOH(aq)$

Plan Your Strategy
Write the complete chemical equation for the reaction.
Write $Na_2SO_4(aq)$, $Sr(OH)_2(aq)$, and $NaOH(aq)$ as ions. Leave $SrSO_4(s)$ as a formula unit since this ionic compound has low solubility.
Write the complete ionic equation for the reaction.
Identify the spectator ions, and cancel them on both sides of the equation.
Write the net ionic equation.

Act on Your Strategy
The precipitate is strontium sulfate, $SrSO_4(s)$.
Skeleton equation:
$Na_2SO_4(aq) + Sr(OH)_2(aq) \rightarrow SrSO_4(s) + NaOH(aq)$
Complete chemical equation:
$Na_2SO_4(aq) + Sr(OH)_2(aq) \rightarrow SrSO_4(s) + 2NaOH(aq)$
Complete ionic equation:
$$\cancel{2Na^{+}(aq)} + SO_4^{2-}(aq) + Sr^{2+} + \cancel{2OH^{-}(aq)} \rightarrow SrSO_4(s) + \cancel{2Na^{+}(aq)} + \cancel{2OH^{-}(aq)}$$
Net ionic equation:
$Sr^{2+}(aq) + SO_4^{2-}(aq) \rightarrow SrSO_4(s)$

Check Your Solution
The net ionic equation is balanced, including the charges on the ions.

3. Practice Problem (page 410)

Write the net ionic equation for this reaction:

$$MgCl_2(aq) + NaOH(aq) \rightarrow Mg(OH)_2(s) + NaCl(aq)$$

What Is Required?

You need to write the net ionic equation for the reaction.

What Is Given?

You know that the reaction between magnesium chloride and sodium hydroxide is a double displacement reaction.
You know that magnesium hydroxide, $Mg(OH)_2(s)$, is the precipitate.
You know the skeleton equation:
$MgCl_2(aq) + NaOH(aq) \rightarrow Mg(OH)_2(s) + NaCl(aq)$

Plan Your Strategy

Write the complete chemical equation for the reaction.
Write $MgCl_2(aq)$, $NaOH(aq)$, and $NaCl(aq)$ as ions. Leave $Mg(OH)_2(s)$ as a formula unit since this ionic compound has low solubility.
Write the complete ionic equation for the reaction.
Identify the spectator ions, and cancel them on both sides of the equation.
Write the net ionic equation.

Act on Your Strategy

The precipitate is magnesium hydroxide, $Mg(OH)_2(s)$.
Skeleton equation:
$MgCl_2(aq) + NaOH(aq) \rightarrow Mg(OH)_2(s) + NaCl(aq)$
Complete chemical equation:
$MgCl_2(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + 2NaCl(aq)$
Complete ionic equation:

$$Mg^{2+}(aq) + \cancel{2Cl^-(aq)} + \cancel{2Na^+(aq)} + 2OH^-(aq) \rightarrow Mg(OH)_2(s) + \cancel{2Na^+(aq)} + \cancel{2Cl^-(aq)}$$

Net ionic equation:
$Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$

Check Your Solution

The net ionic equation is balanced, including the charges on the ions.

4. Practice Problem (page 410)

Barium sulfate, $BaSO_4(s)$, is used in some types of paint as a white pigment and as a filler. Barium sulfate precipitates when an aqueous solution of barium chloride, $BaCl_2(aq)$, is mixed with an aqueous solution of sodium sulfate, $Na_2SO_4(aq)$. Write the complete chemical equation and the net ionic equation for this reaction.

What Is Required?

You need to write the complete chemical equation and the net ionic equation for the reaction.

What Is Given?

You know that the reaction between barium chloride and sodium sulfate is a double displacement reaction.
You know that barium sulfate, $BaSO_4(s)$, is the precipitate.
You know the skeleton equation:
$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + NaCl(aq)$

Plan Your Strategy

Write the skeleton equation and the complete chemical equation for the reaction.
Write $BaCl_2(aq)$, $Na_2SO(aq)$, and $NaCl(aq)$ as ions. Leave $BaSO_4(s)$ as a formula unit since this ionic compound is of low solubility.
Write the complete ionic equation for the reaction.
Identify the spectator ions, and cancel them on both sides of the equation.
Write the net ionic equation.

Act on Your Strategy

The precipitate is barium sulfate, $BaSO_4(s)$.
Skeleton equation:
$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + NaCl(aq)$
Complete chemical equation:
$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$
Complete ionic equation:

$$Ba^{2+}(aq) + \cancel{2Cl^-(aq)} + \cancel{2Na^+(aq)} + SO_4^{\;2-}(aq) \rightarrow BaSO_4(s) + \cancel{2Na^+} + \cancel{2Cl^-(aq)}$$

Net ionic equation:
$Ba^{2+}(aq) + SO_4^{\;2-}(aq) \rightarrow BaSO_4(s)$

Check Your Solution

The net ionic equation is balanced, including the charges on the ions.

5. Practice Problem (page 410)

Identify the precipitate and the spectator ions in the reaction that occurs when an aqueous solution of sodium sulfide is mixed with an aqueous solution of iron(II) sulfate. Write the net ionic equation.

What Is Required?

You need to identify the precipitate and the spectator ions and write the net ionic equation for the reaction.

What Is Given?

You know that the reaction between sodium sulfide and iron(II) sulfate is a double displacement reaction.

Plan Your Strategy

Predict the products that form in this reaction.
Write the chemical formulas for the reactants and products.
Use the solubility guidelines on page 363 to identify the precipitate.
Write the skeleton equation and the complete chemical equation for the reaction.
Write sodium sulfide, iron(II) sulfate, and sodium sulfate as ions. Leave iron(II) sulfide as a formula unit since this ionic compound has low solubility.
Write the complete ionic equation for the reaction.
Identify the spectator ions, and cancel them on both sides of the equation.
Write the net ionic equation.

Act on Your Strategy

The products are sodium sulfide and iron(II) sulfate.
Chemical formulas for the reactants: sodium sulfide, $Na_2S(aq)$; iron(II) sulfate, $FeSO_4(aq)$
Chemical formulas for the products: iron(II) sulfide, $FeS(s)$; sodium sulfate, $Na_2SO_4(aq)$
The precipitate is iron(II) sulfide, $FeS(s)$.
Skeleton equation:
$Na_2S(aq) + FeSO_4(aq) \rightarrow FeS(s) + Na_2SO_4(aq)$
Complete chemical equation:
$Na_2S(aq) + FeSO_4(aq) \rightarrow FeS(s) + Na_2SO_4(aq)$
Complete ionic equation:

$$\cancel{2Na^{+}(aq)} + S^{2-}(aq) + Fe^{2+}(aq) + \cancel{SO_4^{\ 2-}(aq)} \rightarrow FeS(s) + \cancel{2Na^{+}(aq)} + \cancel{SO_4^{\ 2-}(aq)}$$

The spectator ions are $Na^{+}(aq)$ and $SO_4^{\ 2-}(aq)$.
Net ionic equation:
$Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$

Check Your Solution
The net ionic equation is balanced, including the charges on the ions. The spectator ions have been correctly identified.

6. Practice Problem (page 410)
Identify the spectator ions in the reaction between each pair of aqueous solutions. Then write the net ionic equation for each reaction.
a. ammonium phosphate and zinc sulfate
b. lithium carbonate and nitric acid
c. sulfuric acid and barium hydroxide

What Is Required?
For each reaction, you need to identify the spectator ions and write the net ionic equation.

What Is Given?
You know that the reactions are double displacement reactions.

Plan Your Strategy
For each reaction, the strategy is as follows:
Predict the name of the products in these double displacement reactions.
Write the chemical formulas for the reactants and the products.
Use the solubility guidelines on page 363 to identify the precipitate.
Write the skeleton equation and the complete chemical equation.
Write the complete ionic equation.
Identify the spectator ions, and cancel them on both sides of the equation.
Write the net ionic equation.

Act on Your Strategy
a. ammonium phosphate and zinc sulfate
The products are predicted to be zinc phosphate and ammonium sulfate.
Chemical formulas for the reactants: ammonium phosphate, $(NH_4)_3PO_4(aq)$; zinc sulfate, $ZnSO_4(aq)$
Chemical formulas for the products: zinc phosphate, $Zn_3(PO_4)_3(s)$; ammonium sulfate, $(NH_4)_2SO_4(aq)$
The precipitate is zinc phosphate, $Zn_3(PO_4)_3(s)$.
Skeleton equation:
$(NH_4)_3PO_4(aq) + ZnSO_4(aq) \rightarrow Zn_3(PO_4)_2(s) + (NH_4)_2SO_4(aq)$
Complete chemical equation:
$2(NH_4)_3PO_4(aq) + 3ZnSO_4(aq) \rightarrow Zn_3(PO_4)_2(s) + 3(NH_4)_2SO_4(aq)$
Complete ionic equation:

$$\cancel{6NH_4^+(aq)} + 2PO_4^{3-}(aq) + 3Zn^{2+}(aq) + \cancel{3SO_4^{2-}(aq)} \rightarrow Zn_3(PO_4)_2(s) + \cancel{6NH_4^+(aq)} + \cancel{3SO_4^{2-}(aq)}$$

The spectator ions are $NH_4^+(aq)$ and $SO_4^{2-}(aq)$.
Net ionic equation:
$3Zn^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Zn_3(PO_4)_2(s)$

b. lithium carbonate and nitric acid
The products are predicted to be lithium nitrate and carbonic acid.
Chemical formulas for the reactants: lithium carbonate, $Li_2CO_3(aq)$; nitric acid, $HNO_3(aq)$
Chemical formulas for the products: lithium nitrate, $LiNO_3(aq)$; carbonic acid, $H_2CO_3(aq)$
Both products are very soluble, so there is no precipitate.
Skeleton equation:
$Li_2CO_3(aq) + HNO_3(aq) \rightarrow LiNO_3(aq) + H_2CO_3(aq)$
Complete chemical equation:
$Li_2CO_3(aq) + 2HNO_3(aq) \rightarrow 2LiNO_3(aq) + H_2CO_3(aq)$
Complete ionic equation:
In aqueous solution, carbonic acid exists as $H_2O(\ell) + CO_2(g)$.
In aqueous solution, lithium nitrate exists as $Li^+(aq) + NO_3^-(aq)$.

$$\cancel{2Li^+(aq)} + CO_3^{2-}(aq) + 2H^+(aq) + \cancel{2NO_3^-(aq)} \rightarrow \cancel{2Li^+(aq)} + \cancel{2NO_3^-(aq)} + H_2O(l) + CO_2(g)$$

The spectator ions are $Li^+(aq)$ and $NO_3^-(aq)$.
Net ionic equation:
$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow H_2O(\ell) + CO_2(g)$

c. sulfuric acid and barium hydroxide
The products are predicted to be water and barium sulfate.
Chemical formulas for the reactants: sulfuric acid, $H_2SO_4(aq)$; barium hydroxide, $Ba(OH)_2(aq)$
Chemical formulas for the products: water, $H_2O(\ell)$; barium sulfate, $BaSO_4(s)$
The precipitate is barium sulfate, $BaSO_4(s)$.
Skeleton equation:
$H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow H_2O(\ell) + BaSO_4(s)$
Complete chemical equation:
$H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow 2H_2O(\ell) + BaSO_4(s)$
Complete ionic equation: 2
$H^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2OH^-(aq) \rightarrow 2H_2O(\ell) + BaSO_4(s)$
There are no spectator ions.

Check Your Solution
For each reaction, the net ionic equation is balanced, including the charges on the ions.
The spectator ions are correctly identified.

7. Practice Problem (page 410)

When aqueous solutions of sodium iodide and lead(II) nitrate are mixed, a bright yellow precipitate of lead(II) iodide forms. Write a net ionic equation to represent this reaction.

What Is Required?

You need to write the net ionic equation for the reaction.

What Is Given?

You know that the reaction between sodium iodide and lead(II) nitrate is a double displacement reaction.

You know that the precipitate is lead(II) iodide, $PbI_2(s)$.

Plan Your Strategy

Predict the name of the second product that forms.

Write the chemical formulas for the reactants and the products.

Write the skeleton equation and the complete chemical equation for the reaction.

Write the complete ionic equation for the reaction.

Identify the spectator ions, and cancel them on both sides of the equation.

Write the net ionic equation.

Act on Your Strategy

The other product that is predicted to form is sodium nitrate.

Chemical formulas for the reactants: sodium iodide, NaI(aq); lead(II) nitrate, $Pb(NO_3)_2(aq)$

Chemical formulas for the products: lead(II) iodide, $PbI_2(s)$; sodium nitrate, $NaNO_3(aq)$

Skeleton equation:

$NaI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + NaNO_3(aq)$

Complete chemical equation:

$2NaI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2NaNO_3(aq)$

Complete ionic equation:

$$\cancel{2Na^+(aq)} + 2I^-(aq) + Pb^{2+}(aq) + \cancel{2NO_3^-(aq)} \rightarrow PbI_2(s) + \cancel{2Na^+(aq)} + \cancel{2NO_3^-(aq)}$$

The spectator ions are $Na^+(aq)$ and $NO_3^-(aq)$.

Net ionic equation:

$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$

Check Your Solution

The net ionic equation is balanced, including the charges on the ions. The spectator ions have been correctly identified.

8. Practice Problem (page 410)

A chemical reaction can be represented by the following net ionic equation: $2Al^{3+}(aq) + 3Cr_2O_7^{2-}(aq) \rightarrow Al_2(Cr_2O_7)_3(s)$. Suggest two aqueous solutions that could be mixed to cause this reaction.

What Is Required?

You need to suggest two aqueous solutions that, when mixed, will result in the given net ionic equation.

What Is Given?

You know the net ionic equation: $2Al^{3+}(aq) + 3Cr_2O_7^{2-}(aq) \rightarrow Al_2(Cr_2O_7)_3(s)$

Plan Your Strategy

You need to start with reactants that have the cation $Al^{3+}(aq)$ and the anion $Cr_2O_7^{2-}(aq)$.
Refer to the solubility guidelines and Table 8.3 on page 363, and select a soluble compound of $Al^{3+}(aq)$ and $Cr_2O_7^{2-}(aq)$.
Write the skeleton equation and the complete chemical equation for the reaction.
Write the complete ionic equation for the reaction.
Identify the spectator ions, and cancel them on both sides of the equation.
Write the net ionic equation and confirm that it is the same as the given net ionic equation.

Act on Your Strategy

Use the reactants aluminum nitrate, $Al(NO_3)_3(aq)$, and potassium dichromate, $K_2Cr_2O_7(aq)$.
Skeleton equation:
$Al(NO_3)_3(aq) + K_2Cr_2O_7(aq) \rightarrow Al_2(Cr_2O_7)_3(s) + KNO_3(aq)$
Complete chemical equation:
$2Al(NO_3)_3(aq) + 3K_2Cr_2O_7(aq) \rightarrow Al_2(Cr_2O_7)_3(s) + 6KNO_3(aq)$
Complete ionic equation:

$$2Al^{3+}(aq) + \cancel{6NO_3^-(aq)} + \cancel{6K^+(aq)} + 3Cr_2O_7^{2-}(aq) \rightarrow Al_2(Cr_2O_7)_3(s) + \cancel{6NO_3^-(aq)} + \cancel{6K^+(aq)}$$

The spectator ions are $K^+(aq)$ and $NO_3^-(aq)$.
Net ionic equation:
$2Al^{3+}(aq) + 3Cr_2O_7^{2-}(aq) \rightarrow Al_2(Cr_2O_7)_3(s)$

Check Your Solution

The net ionic equation is balanced, including the charges on the ions, and there is agreement with the given net ionic equation.

9. Practice Problem (page 410)

Iron(III) ions, Fe^{3+}(aq), can be precipitated from a solution by adding potassium hydroxide, KOH(aq). Write the net ionic equation for the reaction between iron(III) nitrate, $Fe(NO_3)_3$(aq), and potassium hydroxide. Identify the spectator ions.

What Is Required?

You need to write the net ionic equation for the reaction between the Fe^{3+}(aq) ion and OH^-(aq) ion and identify the spectator ions.

What Is Given?

You know the reactants: iron(III) nitrate, $Fe(NO_3)_3$(aq); potassium hydroxide, KOH(aq)

Plan Your Strategy

The reaction between potassium hydroxide and iron(III) nitrate is a double displacement reaction.
Predict the products that form in this reaction.
Write the skeleton equation and the complete chemical equation for the reaction.
Write the complete ionic equation for the reaction.
Identify the spectator ions, and cancel them on both sides of the equation.
Write the net ionic equation.

Act on Your Strategy

The products are predicted to be iron(III) hydroxide and potassium nitrate.
The precipitate is iron(III) hydroxide, $Fe(OH)_3$(s).
Skeleton equation:
$Fe(NO_3)_3(aq) + KOH(aq) \rightarrow Fe(OH)_3(s) + KNO_3(aq)$
Complete chemical equation:
$Fe(NO_3)_3(aq) + 3KOH(aq) \rightarrow Fe(OH)_3(s) + 3KNO_3(aq)$
Complete ionic equation:

$$Fe^{3+}(aq) + \cancel{3NO_3^-(aq)} + \cancel{3K^+(aq)} + 3OH^-(aq) \rightarrow Fe(OH)_3(s) + \cancel{3K^+(aq)} + \cancel{3NO_3^-(aq)}$$

The spectator ions are K^+(aq) and NO_3^-(aq).
Net ionic equation: $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

Check Your Solution

The net ionic equation is balanced, including the charges on the ions, and the spectator ions have been correctly identified.

10. Practice Problem (page 410)

Complete and balance each equation. Then write the corresponding net ionic equation.

a. $Pb(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow$

b. $Co(CH_3COO)_2(aq) + (NH_4)_2S(aq) \rightarrow$

What Is Required?

You need to complete and balance each equation and write the net ionic equation for each reaction.

What Is Given?

You know the reactants in each reaction.
You know that each reaction is a double displacement reaction.

Plan Your Strategy

For each reaction, the strategy is as follows::
Predict the name of the products that form and write the chemical formulas for the products of these double displacement reactions.
Use the solubility guidelines on page 363 to identify the precipitate.
Write the skeleton equation and the complete chemical equation for each reaction.
Write the complete ionic equation for the reactions.
Identify the spectator ions, and cancel them on both sides of both equations.
Write the net ionic equation for each reaction.

Act on Your Strategy

a. $Pb(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow$
The products are predicted to be lead(II) carbonate and sodium nitrate.
The precipitate is lead(II) carbonate, $PbCO_3(s)$.
Skeleton equation:
$Pb(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow PbCO_3(s) + NaNO_3(aq)$
Complete chemical equation:
$Pb(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow PbCO_3(s) + 2NaNO_3(aq)$
Complete ionic equation:

$$Pb^{2+}(aq) + \cancel{2NO_3^-(aq)} + \cancel{2Na^+(aq)} + CO_3^{2-}(aq) \rightarrow PbCO_3(s) + \cancel{2Na^+(aq)} + \cancel{2NO_3^-(aq)}$$

The spectator ions are $Na^+(aq)$ and $NO_3^-(aq)$.
Net ionic equation:
$Pb^{2+}(aq) + CO_3^{2-}(aq) \rightarrow PbCO_3(s)$

b. $Co(CH_3COO)_2(aq) + (NH_4)_2S(aq) \rightarrow$
The products are predicted to be cobalt sulfide and ammonium acetate.
The precipitate is cobalt(II) sulfide, $CoS(s)$.

Skeleton equation:
$Co(CH_3COO)_2(aq) + (NH_4)_2S(aq) \rightarrow CoS(s) + NH_4CH_3COO(aq)$
Complete chemical equation:
$Co(CH_3COO)_2(aq) + (NH_4)_2S(aq) \rightarrow CoS(s) + 2NH_4CH_3COO(aq)$
Complete ionic equation:

$$Co^{2+}(aq) + \cancel{2CH_3COO^-(aq)} + \cancel{2NH_4^+(aq)} + S^{2-}(aq) \rightarrow CoS(s) + \cancel{2NH_4^+(aq)} + \cancel{2CH_3COO^-(aq)}$$

The spectator ions are $NH_4^+(aq)$ and $CH_3COO^-(aq)$.
Net ionic equation:
$Co^{2+}(aq) + S^{2-}(aq) \rightarrow CoS(s)$

Check Your Solution
The net ionic equation is balanced, including the charges on the ions.

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Solutions for Selected Review Questions
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2. Review Question (page 414)
Identify the spectator ions in each reaction.
a. $3CuCl_2(aq) + 2(NH_4)_3PO_4(aq) \rightarrow Cu_3(PO_4)_2(s) + 6NH_4Cl(aq)$
b. $2Al(NO_3)_3(aq) + 3Ba(OH)_2(aq) \rightarrow 2Al(OH)_3(s) + 3Ba(NO_3)_2(aq)$
c. $2NaOH(aq) + MgCl_2(aq) \rightarrow 2NaCl(aq) + Mg(OH)_2(s)$

What Is Required?
For each reaction, you need to identify the spectator ions.

What Is Given?
You know the balanced equation and the product that is the precipitate.

Plan Your Strategy
Write the complete ionic equation for the reaction.
Identify the spectator ions, and cancel them on both sides of the equation.

Act on Your Strategy
a. $3CuCl_2(aq) + 2(NH_4)_3PO_4(aq) \rightarrow Cu_3(PO_4)_2(s) + 6NH_4Cl(aq)$:
Balanced chemical equation:
$3CuCl_2(aq) + 2(NH_4)_3PO_4(aq) \rightarrow Cu_3(PO_4)_2(s) + 6NH_4Cl(aq)$
Complete ionic equation:

$$3Cu^{2+}(aq) + \cancel{6Cl^-(aq)} + \cancel{6NH_4^+(aq)} + 2PO_4^{3-}(aq) \rightarrow Cu_3(PO_4)_2(s) + \cancel{6NH_4^+(aq)} + \cancel{6Cl^-(aq)}$$

The spectator ions are $NH_4^+(aq)$ and $Cl^-(aq)$.

b. $2Al(NO_3)_3(aq) + 3Ba(OH)_2(aq) \rightarrow 2Al(OH)_3(s) + 3Ba(NO_3)_2(aq)$:
Balanced chemical equation:
$2Al(NO_3)_3(aq) + 3Ba(OH)_2(aq) \rightarrow 2Al(OH)_3(s) + 3Ba(NO_3)_2(aq)$
Complete ionic equation:

$$2Al^{3+}(aq) + \cancel{6NO_3^-(aq)} + \cancel{3Ba^{2+}(aq)} + 6OH^-(aq) \rightarrow 2Al(OH)_3(s) + \cancel{3Ba^{2+}(aq)} + \cancel{6NO_3^-(aq)}$$

The spectator ions are $Ba^{2+}(aq)$ and $NO_3^-(aq)$.

c. $2NaOH(aq) + MgCl_2(aq) \rightarrow 2NaCl(aq) + Mg(OH)_2(s)$:
Balanced chemical equation:
$2NaOH(aq) + MgCl_2(aq) \rightarrow 2NaCl(aq) + Mg(OH)_2(s)$
Complete ionic equation:

$$\cancel{2Na^+(aq)} + 2OH^-(aq) + Mg^{2+}(aq) + \cancel{2Cl^-(aq)} \rightarrow \cancel{2Na^+(aq)} + \cancel{2Cl^-(aq)} + Mg(OH)_2(s)$$

The spectator ions are $Na^+(aq)$ and $Cl^-(aq)$.

Check Your Solution
The complete ionic equation is balanced, including the charges on the ions, and the spectator ions are correctly identified.

3. Review Question (page 414)
Write a net ionic equation for each reaction in question 2.

What Is Required?
You need to write the net ionic equations.

What Is Given?
From question 2, you know the complete ionic equations and the spectator ions.

Plan Your Strategy
Cancel the spectator ions from the complete ionic equations that are given in question 2.
Write the net ionic equation.

Act on Your Strategy
a. $3Cu^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Cu_3(PO_4)_2(s)$
b. $2Al^{3+}(aq) + 6OH^-(aq) \rightarrow 2Al(OH)_3(s)$
c. $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$

Check Your Solution
The net ionic equation is balanced, including the charges on the ions.

4. Review Question (page 414)

An aqueous solution of copper(II) sulfate is mixed with an aqueous solution of sodium carbonate.

a. State the name and formula for the precipitate that forms.
b. Write the net ionic equation for the reaction.
c. Identify the spectator ions.

What Is Required?

You need to write the name and formula for the precipitate, identify the spectator ions, and write the net ionic equation for the reaction between aqueous copper(II) sulfate and aqueous sodium carbonate.

What Is Given?

You know the reactants are aqueous copper(II) sulfate and aqueous sodium carbonate. You know that the reaction is a double displacement reaction.

Plan Your Strategy

Write the chemical formula for the reactants and products of this double displacement reaction.
Refer to the solubility guidelines on page 363 to identify the precipitate.
Write the skeleton equation and the complete chemical equation for the reaction.
Write the complete ionic equation for the reaction.
Identify the spectator ions, and cancel them on both sides of the equation.
Write the net ionic equation.

Act on Your Strategy

Chemical formulas for the reactants: copper(II) sulfate, $CuSO_4(aq)$; sodium carbonate, $Na_2CO_3(aq)$
Chemical formulas for the products: copper(II) carbonate, $CuCO_3(s)$; sodium sulfate, $Na_2SO_4(aq)$

a. name and formula for the precipitate
The precipitate is copper(II) carbonate.
The formula for the precipitate is $CuCO_3(s)$.
Skeleton equation:
$CuSO_4(aq) + Na_2CO_3(aq) \rightarrow CuCO_3(s) + Na_2SO_4(aq)$
Balanced equation:
$CuSO_4(aq) + Na_2CO_3(aq) \rightarrow CuCO_3(s) + Na_2SO_4(aq)$
Complete ionic equation:

$$Cu^{2+}(aq) + \cancel{SO_4^{2-}(aq)} + \cancel{2Na^+(aq)} + CO_3^{2-}(aq) \rightarrow CuCO_3(s) + \cancel{2Na^+(aq)} + \cancel{SO_4^{2-}(aq)}$$

b. net ionic equation
$Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$

c. spectator ions
$Na^{+}(aq)$ and $SO_4^{2-}(aq)$

Check Your Solution
The net ionic equation is balanced, including the charges on the ions, and the spectator ions have been correctly identified.

5. Review Question (page 414)
For each of the following net ionic equations, list two soluble ionic compounds that can be mixed together in solution to produce the reaction represented by the equation. (**Note:** There are many correct answers.)
a. $3Ba^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ba_3(PO_4)_2(s)$
b. $Mg^{2+}(aq) + 2OH^{-}(aq) \rightarrow Mg(OH)_2(s)$
c. $2Al^{3+}(aq) + 3Cr_2O_7^{2-}(aq) \rightarrow Al_2(Cr_2O_7)_3(s)$

What Is Required?
You need to list soluble reactants that, when mixed, will result in the given net ionic equations.

What Is Given?
You know the balanced net ionic equation for the reactions.

Plan Your Strategy
Refer to the solubility guidelines on page 363 and select compounds that are soluble for the cations and anions that are the reactants in each net ionic equation.

Act on Your Strategy
a. $3Ba^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ba_3(PO_4)_2(s)$
Barium chloride:
$BaCl_2(s) \rightarrow Ba^{2+}(aq) + 2Cl^{-}(aq)$;
Sodium phosphate:
$Na_3PO_4(s) \rightarrow 3Na^{+}(aq) + PO_4^{3-}(aq)$

b. $Mg^{2+}(aq) + 2OH^{-}(aq) \rightarrow Mg(OH)_2(s)$
Magnesium nitrate:
$Mg(NO_3)_2(s) \rightarrow Mg^{2+}(aq) + 2NO_3^{-}(aq)$
Potassium hydroxide:
$KOH(s) \rightarrow K^{+}(aq) + OH^{-}(aq)$

c. $2Al^{3+}(aq) + 3Cr_2O_7^{2-}(aq) \rightarrow Al_2(Cr_2O_7)_3(s)$
Aluminum sulfate:
$Al_2(SO_4)_3(s) \rightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$
Sodium dichromate:
$Na_2Cr_2O_7(s) \rightarrow 2Na^{+}(aq) + Cr_2O_7^{2-}(aq)$

Check Your Solution
The compounds are all soluble and the ions have the correct charge.

6. Review Question (page 414)
Explain why there are many correct answers for question 5.

There are many soluble compounds that have the required cations and anions.

10. Review Question (page 414)
Lithium carbonate is the active ingredient in some anti-depression medications. What tests could you perform to confirm the presence of lithium carbonate, $Li_2CO_3(s)$, in a tablet?

A flame test giving a crimson red colour indicates the presence of the lithium cation, $Li^{+}(aq)$. Dilute hydrochloric acid added to the compound producing vigorous bubbling of carbon dioxide indicates the presence of the carbonate anion, $CO_3^{2-}(aq)$.

11. Review Question (page 414)
Limewater is a solution of calcium hydroxide, $Ca(OH)_2(aq)$. It can be used to test for the presence of carbon dioxide. When carbon dioxide is bubbled through limewater, a milky-white precipitate is produced.
a. Write a chemical equation and a net ionic equation to show what happens when carbon dioxide is bubbled through limewater.
b. Is this test an example of qualitative or quantitative analysis? Explain your answer.

a. chemical equation and net ionic equation
$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(\ell)$
$Ca^{2+}(aq) + 2OH^{-}(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(\ell)$
b. type of analysis
The tests are qualitative because they identify what substance is present, but not how much is present.

12. Review Question (page 414)

An ion in a solution forms a yellow precipitate when sodium iodide, NaI(aq), is added to the solution. The precipitate produces a blue-white colour when it is heated in a flame.
a. Suggest a formula for the ion and a formula for the precipitated compound.
b. Write a net ionic equation to represent the reaction.

a. The blue-white flame colour indicates that the lead(II) cation, Pb^{2+}(aq), is present. In aqueous solution, sodium iodide, NaI(aq) exists as Na^{+}(aq) and I^{-}(aq). The Pb^{2+}(aq) will react with the iodide ion to form lead(II) iodide, PbI_2(s), which is a yellow precipitate.
b. $Pb^{2+}(aq) + 2I^{-}(aq) \rightarrow PbI_2(s)$

14. Review Question (page 414)

To answer the following questions, refer to the solubility guidelines in Section 8.2 (see page 363).
a. What aqueous solution will precipitate Pb^{2+}(aq) ions but not Cu^{+}(aq) or Mg^{2+}(aq) ions?
b. What aqueous solution will precipitate Cu^{+}(aq) ions but not $Mg^{2+(}$aq) ions?
c. Using the relationship your answers to parts **a** and **b**, outline a procedure that would allow you to precipitate the Pb^{2+}(aq) ions, followed by Cu^{+}(aq) ions and then Mg^{2+}(aq) ions.

a. A solution of sodium sulfate, Na_2SO_4(aq), will precipitate Pb^{2+}(aq) as $PbSO_4$(s).
b. A solution containing a halide such as sodium chloride, NaCl(aq); sodium bromide, NaBr(aq); or sodium iodide, NaI(aq) will precipitate Cu^{+}(aq) as CuCl(s), CuBr(s), or CuI(s).
c. Procedure:

- add Na_2SO_4(aq)
- filter
- to filtrate, add NaCl(aq)
- filter
- to filtrate, add KOH(aq)

Pb^{2+}(aq)	Na_2SO_4 (aq)	Cu^{+}(aq)	NaCl(aq)	Mg^{2+}(aq)	KOH(aq)
Cu^{+}(aq)	————>	Mg^{2+}(aq)	————>		————>
Mg^{2+}(aq)	↓		↓		↓
	$PbSO_4$(s)		CuCl(s)		

$Mg(OH)_2$(s)

Section 9.2 Solution Stoichiometry

Solutions for Practice Problems

Student Edition page 417

11. Practice Problem (page 417)

If 8.5 g of pure ammonium phosphate, $(NH_4)_3PO_4(s)$, is dissolved in distilled water to make 400 mL of solution, what are the concentrations (in moles per litre) of the ions in the solution?

What Is Required?

You need to find the molar concentration, *c*, of the ions in a solution of ammonium phosphate.

What Is Given?

You know the volume of the ammonium phosphate solution: 400 mL
You know the mass of ammonium phosphate, $(NH_4)_3PO_4(s)$: 8.5 g

Plan Your Strategy

Write the balanced chemical equation for the dissolution of ammonium phosphate, $(NH_4)_3PO_4(s)$.
Determine the molar mass of $(NH_4)_3PO_4(s)$.

Calculate the amount in moles of $(NH_4)_3PO_4(s)$ using the relationship $n = \frac{n}{M}$.

Convert the volume from millilitres to litres: 1 mL = 1×10^{-3} L

Calculate the concentration of $(NH_4)_3PO_4(aq)$ using the relationship $c = \frac{n}{V}$.

Equate the mole ratios and cross multiply to solve for *n*, the amount in moles of $(NH_4)_3PO_4(s)$.

Act on Your Strategy

Balanced equation:	$(NH_4)_3PO_4(s) \rightarrow$	$3NH_4^+(aq) +$	$PO_4^{3-}(aq)$
Mole ratio:	1 mole	3 moles	1 mole

Molar mass, *M*, of $(NH_4)_3PO_4(s)$:

$$\begin{aligned} M_{(NH_4)_3PO_4} &= 3M_N + 12M_H + 1M_P + 4M_O \\ &= 3(14.01\text{g/mol}) + 12(1.01 \text{ g/mol}) + 1(30.97 \text{ g/mol}) + 4(16.00 \text{ g/mol}) \\ &= 149.12 \text{ g/mol} \end{aligned}$$

Amount in moles, *n*, of $(NH_4)_3PO_4(s)$:

$$n_{(NH_4)_3PO_4} = \frac{m}{M}$$
$$= \frac{8.5\ \cancel{g}}{149.12\ \cancel{g}/\text{mol}}$$
$$= 5.700 \times 10^{-2}\ \text{mol}$$

Amount in moles, *n*, of $NH_4^+(aq)$:

$$\frac{1\ \text{mol}\ (NH_4)_3PO_4}{3\ \text{mol}\ NH_4^+} = \frac{0.005700\ \text{mol}\ (NH_4)_3PO_4}{n_{NH_4^+}}$$
$$n_{NH_4^+} = \frac{3\ \text{mol}\ NH_4^+ \times 0.005700\ \cancel{\text{mol}\ (NH_4)_3PO_4}}{1\ \cancel{\text{mol}\ (NH_4)_3PO_4}}$$
$$= 0.0171\ \text{mol}$$

Amount in moles, *n*, of $PO_4^{3-}(aq)$:

$$\frac{1\ \text{mol}\ (NH_4)_3PO_4}{1\ \text{mol}\ PO_4^{3-}} = \frac{0.005700\ \text{mol}\ (NH_4)_3PO_4}{n_{PO_4^{3-}}}$$
$$n_{PO_4^{3-}} = \frac{1\ \text{mol}\ PO_4^{3-} \times 0.005700\ \cancel{\text{mol}\ (NH_4)_3PO_4}}{1\ \cancel{\text{mol}\ (NH_4)_3PO_4}}$$
$$= 0.005700\ \text{mol}$$

Volume of solution:

$$V = 400\ \cancel{\text{mL}} \times 1 \times 10^{-3}\ \text{L}/\cancel{\text{mL}}$$
$$= 0.400\ \text{L}$$

Concentration of $NH_4^+(aq)$:

$$c = \frac{n}{V}$$
$$= \frac{1.71 \times 10^{-2}\ \text{mol}}{0.400\ \text{L}}$$
$$= 0.04275\ \text{mol/L}$$
$$= 0.04\ \text{mol/L}$$

Concentration of $PO_4^{3-}(aq)$:

$$
\begin{aligned}
c &= \frac{n}{V} \\
&= \frac{0.005700 \times 10^{-2}\ \text{mol}}{0.400\ \text{L}} \\
&= 0.01425\ \text{mol/L} \\
&= 0.01\ \text{mol/L}
\end{aligned}
$$

The concentration of the ammonium ion, $NH_4^+(aq)$, is 0.04 mol/L.
The concentration of the phosphate ion, $PO_4^{3-}(aq)$, is 0.01 mol/L.

Check Your Solution
The units for amount and concentration are correct. The answer has one significant digit and seems reasonable.

12. Practice Problem (page 417)
A strip of zinc metal was placed in a beaker that contained 120 mL of a solution of copper(II) nitrate, $Cu(NO_3)_2(aq)$. The mass of the copper produced was 0.813 g. Find the initial concentration of the copper(II) nitrate solution.

What Is Required?
You need to find the molar concentration, *c*, of the copper(II) nitrate solution.

What Is Given?
You know the volume of the copper(II) nitrate solution: 120 mL
You know the mass of copper precipitated: 0.813 g
You know the other reactant: zinc

Plan Your Strategy
Write the chemical equation for the single displacement reaction.
Use the periodic table to determine the atomic molar mass of Cu(s).
Calculate the amount in moles of Cu(s) using the relationship $n = \frac{n}{M}$.
Calculate the amount in moles of copper(II) nitrate using the mole ratio in the balanced equation.
Convert the volume from millilitres to litres: 1 mL = 1×10^{-3} L
Calculate the concentration of copper(II) nitrate using the relationship $c = \frac{n}{V}$.

Act on Your Strategy

Balanced equation: $Zn(s) + Cu(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + Cu(s)$

Mole ratio: 1 mole 1 mole 1 mole 1 mole

Molar mass, *M*, of Cu(s)

$M_{Cu} = 63.55$ g/mol (from the periodic table)

Amount in moles, *n*, of Cu(s):

$$\begin{aligned} n_{Cu} &= \frac{m}{M} \\ &= \frac{0.813\ \cancel{g}}{63.55\ \cancel{g}/\text{mol}} \\ &= 0.01279\ \text{mol} \end{aligned}$$

Amount in moles, *n*, of $Cu(NO_3)_2(aq)$:

$$\frac{1\ \text{mol Cu(NO}_3)_2}{1\ \text{mol Cu}} = \frac{n_{Cu(NO_3)_2}}{0.01279\ \text{mol Cu}}$$

$$\begin{aligned} n_{Cu(NO_3)_2} &= \frac{1\ \text{mol Cu(NO}_3)_2 \times 0.01279\ \cancel{\text{mol Cu}}}{1\ \cancel{\text{mol Cu}}} \\ &= 0.01279\ \text{mol} \end{aligned}$$

Volume of solution:

$$\begin{aligned} V &= 120\ \cancel{\text{mL}} \times 1 \times 10^{-3}\ \text{L}/\cancel{\text{mL}} \\ &= 0.120\ \text{L} \end{aligned}$$

Concentration of copper(II) nitrate solution:

$$\begin{aligned} c &= \frac{n}{V} \\ &= \frac{0.01279\ \text{mol}}{0.120\ \text{L}} \\ &= 0.10658\ \text{mol/L} \\ &= 0.11\ \text{mol/L} \end{aligned}$$

The initial concentration of the copper(II) nitrate solution was 0.11 mol/L.

Check Your Solution

The units for amount and concentration are correct. The answer has two significant digits and seems reasonable.

13. Practice Problem (page 417)

When 75.0 mL of silver nitrate, $AgNO_3$(aq), was treated with excess ammonium carbonate, $(NH_4)_2CO_3$(aq), 2.47 g of dry precipitate was recovered. Write the net ionic equation for the reaction, and calculate the concentration of the original silver nitrate solution.

What Is Required?

You need to write the net ionic equation for the reaction and determine the initial concentration of the silver nitrate solution.

What Is Given?

You know the volume of the silver nitrate solution: 75 mL
You know the mass of precipitate: 2.47 g
You know the other reactant: ammonium carbonate

Plan Your Strategy

Predict the products that form in this double displacement reaction.
Write the balanced equation for the double displacement reaction.
Refer to the solubility guidelines on page 363 to determine the name of the precipitate.
Write the complete ionic equation for the reaction.
Write the net ionic equation for the reaction.
Determine the molar mass of the precipitate.

Calculate the amount in moles of the precipitate using the relationship $n = \frac{n}{M}$.

Equate the mole ratios and solve for the amount in moles of $AgNO_3$(aq).
Convert the volume from millilitres to litres: 1 mL = 1×10^{-3} L

Calculate the concentration of $AgNO_3$(aq) using the relationship $c = \frac{n}{V}$.

Act on Your Strategy

The products are predicted to be ammonium nitrate and silver carbonate.

Balanced equation:	$2AgNO_3(aq)$ +	$(NH_4)_2CO_3(aq)$ →	$2NH_4NO_3(aq)$ +	$Ag_2CO_3(s)$
Mole ratio:	2 moles	1 mole	2 moles	1 mole

Complete ionic equation:

$$2Ag^+(aq) + \cancel{2NO_3^-(aq)} + \cancel{2NH_4^+(aq)} + CO_3^{2-}(aq) \rightarrow \cancel{2NH_4^+(aq)} + \cancel{2NO_3^-(aq)} + Ag_2CO_3(s)$$

Net ionic equation:

$$2Ag^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3(s)$$

Molar mass, *M*, of the precipitate, $Ag_2CO_3(s)$:

$$M_{Ag_2CO_3} = 2M_{Ag} + 1M_{c} + 3M_{O}$$
$$= 2(107.87 \text{ g/mol}) + 1(12.01 \text{ g/mol}) + 3(16.00 \text{ g/mol})$$
$$= 275.75 \text{ g/mol}$$

Amount in moles, *n*, of precipitate, $Ag_2CO_3(s)$:

$$n_{Ag_2CO_3} = \frac{m}{M}$$
$$= \frac{2.47\ \cancel{\text{g}}}{275.75\ \cancel{\text{g}}\text{/mol}}$$
$$= 8.9573 \times 10^{-3} \text{ mol}$$

Amount in moles of $AgNO_3(aq)$:

$$\frac{2 \text{ mol AgNO}_3}{1 \text{ mol Ag}_2\text{CO}_3} = \frac{n_{AgNO_3}}{8.9573 \times 10^{-3} \text{ mol Ag}_2\text{CO}_3}$$
$$n_{AgNO_3} = \frac{2 \text{ mol AgNO}_3 \times 8.9573 \times 10^{-3}\ \cancel{\text{mol Ag}_2\text{CO}_3}}{1\ \cancel{\text{mol Ag}_2\text{CO}_3}}$$
$$= 1.7914 \times 10^{-2} \text{ mol}$$

Volume of solution:

$$V = 75\ \cancel{\text{mL}} \times 1 \times 10^{-3} \text{ L/}\cancel{\text{mL}}$$
$$= 0.075 \text{ L}$$

Concentration of silver nitrate solution:

$$c = \frac{n}{V}$$
$$= \frac{1.7914 \times 10^{-2} \text{ mol}}{0.075 \text{ L}}$$
$$= 2.3885 \times 10^{-1} \text{ mol/L}$$
$$= 2.39 \times 10^{-1} \text{ mol/L}$$

The concentration of the original silver nitrate solution was 2.39×10^{-1} mol/L.

Check Your Solution

The units for amount and concentration are correct. The answer has three significant digits and seems reasonable.

14. Practice Problem (page 417)

When an excess of sodium sulfide, $Na_2S(aq)$, was added to 125 mL of 0.100 mol/L iron(II) nitrate, $Fe(NO_3)_2(aq)$, a black precipitate formed. Identify the precipitate, and calculate the maximum mass of precipitate that can be collected from the reaction.

What Is Required?

You need to name and calculate the maximum mass of precipitate that is produced in a reaction.

What Is Given?

You know the volume of the iron(II) nitrate solution: 125 mL
You know the initial concentration of iron(II) nitrate: 0.100 mol/L
You know the other reactant: aqueous sodium sulfide, $Na_2S(aq)$

Plan Your Strategy

Predict the products that form in this double displacement reaction.
Refer to the solubility guidelines on page 363 to identify the precipitate.
Write the balanced equation for the reaction.
Convert the volume from millilitres to litres: 1 mL = 1 × 10^{-3} L
Calculate the amount in moles of iron(II) nitrate using the relationship $n = c \times V$.
Equate the mole ratios and solve for the amount in moles of precipitate.
Use the periodic table to determine the molar mass of the precipitate.
Calculate the mass of the precipitate using the relationship $m = n \times M$.

Act on Your Strategy

The precipitate is iron(II) sulfide, FeS(s).

Balanced equation:	$Na_2S(aq)$ +	$Fe(NO_3)_2(aq)$ →	$2NaNO_3(aq)$ +	$FeS(s)$
Mole ratio:	1 mole	1 mole	2 moles	1 mole

Volume of solution:

$$V = 125\ \cancel{\text{mL}} \times 1 \times 10^{-3}\ \text{L}/\cancel{\text{mL}}$$
$$= 0.125\ \text{L}$$

Amount in moles, *n*, of $Fe(NO_3)_2(aq)$:

$$n_{Fe(NO_3)_2} = c \times V$$
$$= 0.100\ \text{mol}/\cancel{\text{L}} \times 0.125\ \cancel{\text{L}}$$
$$= 0.0125\ \text{mol}$$

Amount in moles, *n*, of the precipitate, FeS(s):

$$\frac{1\ \text{mol FeS}}{1\ \text{mol Fe(NO}_3)_2} = \frac{n_{\text{FeS}}}{0.0125\ \text{mol Fe(NO}_3)_2}$$

$$\begin{aligned} n_{\text{FeS}} &= \frac{1\ \text{mol FeS} \times 0.0125\ \cancel{\text{mol Fe(NO}_3)_2}}{1\ \cancel{\text{mol Fe(NO}_3)_2}} \\ &= 0.0125\ \text{mol} \end{aligned}$$

Molar mass, *M*, of the precipitate, FeS(s):

$$\begin{aligned} M_{\text{FeS}} &= 1M_{\text{Fe}} + 1M_{\text{S}} \\ &= 1(55.85\ \text{g/mol}) + 1(32.07\ \text{g/mol}) \\ &= 87.92\ \text{g/mol} \end{aligned}$$

Mass, *m*, of FeS(s):

$$\begin{aligned} m_{\text{FeS}} &= n \times M \\ &= 0.0125\ \cancel{\text{mol}} \times 87.92\ \text{g}/\cancel{\text{mol}} \\ &= 1.099\ \text{g} \\ &= 1.10\ \text{g} \end{aligned}$$

The precipitate is iron(II) sulfide, FeS(s), and the maximum mass that can be collected from the reaction is 1.10 g.

Check Your Solution
The units for amount and concentration are correct. The answer has three significant digits and seems reasonable.

15. Practice Problem (page 417)
What mass of silver chloride, AgCl(s), can be precipitated from 75 mL of 0.25 mol/L silver nitrate, $AgNO_3$(aq), by adding excess magnesium chloride, $MgCl_2$(aq)?

What Is Required?
You need to calculate the mass of silver chloride that will precipitate in a reaction.

What Is Given?
You know the volume of the silver nitrate solution: 75 mL
You know the initial concentration of silver nitrate: 0.25 mol/L
You know the other reactant is aqueous magnesium chloride, $MgCl_2$(aq).

Plan Your Strategy

Predict the other product that forms in this double displacement reaction.
Write the balanced equation for the reaction.
Convert the volume to litres: 1 mL = 1×10^{-3} L
Calculate the amount in moles of silver nitrate using the relationship $n = c \times V$.
Equate the mole ratios and solve for the amount in moles of precipitate.
Use the periodic table to determine the molar mass of the precipitate.
Calculate the mass of precipitate using the relationship $m = n \times M$.

Act on Your Strategy

The other product is magnesium nitrate, $Mg(NO_3)_2(aq)$.

Balanced equation:	$MgCl_2(aq)$ +	$2AgNO_3(aq)$ →	$Mg(NO_3)_2(aq)$ +	$2AgCl(s)$
Mole ratio:	1 mole	2 moles	1 mole	2 moles

Volume of solution:

$$\begin{aligned} V &= 75\ \cancel{\text{mL}} \times 1 \times 10^{-3}\ \text{L}/\cancel{\text{mL}} \\ &= 0.075\ \text{L} \end{aligned}$$

Amount in moles, *n*, of $AgNO_3(aq)$:

$$\begin{aligned} n_{\text{AgNO}_3} &= c \times V \\ &= 0.25\ \text{mol}/\cancel{\text{L}} \times 0.075\ \cancel{\text{L}} \\ &= 0.01875\ \text{mol} \end{aligned}$$

Amount in moles, *n*, of precipitate, $AgCl(s)$:

$$\frac{2\ \text{mol AgCl}}{2\ \text{mol AgNO}_3} = \frac{n_{\text{AgCl}}}{0.01875\ \text{mol AgNO}_3}$$

$$\begin{aligned} n_{\text{AgCl}} &= \frac{2\ \text{mol AgCl} \times 0.01875\ \cancel{\text{mol AgNO}_3}}{2\ \cancel{\text{mol AgNO}_3}} \\ &= 0.01875\ \text{mol} \end{aligned}$$

Molar mass, *M*, of the precipitate, $AgCl(s)$:

$$\begin{aligned} M_{\text{AgCl}} &= 1M_{\text{Ag}} + 1M_{\text{Cl}} \\ &= 1(107.87\ \text{g/mol}) + 1(35.45\ \text{g/mol}) \\ &= 143.32\ \text{g/mol} \end{aligned}$$

Mass, *m*, of AgCl(s):

$$m_{AgCl} = n \times M$$
$$= 0.01875\ \cancel{mol} \times 143.32\ g/\cancel{mol}$$
$$= 2.687\ g$$
$$= 2.7\ g$$

The mass of silver chloride that can be precipitated is 2.7 g.

Check Your Solution
The units for amount and concentration are correct. The answer has two significant digits and seems reasonable.

16. Practice Problem (page 417)
What mass of bromine gas can be collected by bubbling excess chlorine gas through 850 mL of a 0.350 mol/L solution of sodium bromide, NaBr(aq)?

What Is Required?
You need to calculate the mass of bromine gas that will collect in a reaction.

What Is Given?
You know the volume of the sodium bromide solution: 850 mL
You know the initial concentration of the sodium bromide solution: 0.350 mol/L
You know the other reactant: chlorine gas, Cl_2(aq)

Plan Your Strategy
Predict the name and formula for the other product that forms in this single displacement reaction.
Write the balanced equation for the reaction.
Convert the volume from millilitres to litres: 1 mL = 1×10^{-3} L
Calculate the amount in moles of sodium bromide solution using the relationship $n = c \times V$.
Equate the mole ratios and solve for the amount in moles of bromine gas.
Use the periodic table to determine the molar mass of the bromine gas, Br_2(g).
Calculate the mass of bromine gas using the relationship $m = n \times M$.

Act on Your Strategy
The other product that forms in this single displacement reaction is sodium chloride, NaCl(aq).

Balanced equation:	Cl_2(g) +	2NaBr(aq) →	2NaCl (aq) +	Br_2(g)
Mole ratio:	1 mole	2 moles	2 moles	1 mole

Volume of solution:

$$V = 850\ \cancel{\text{mL}} \times 1 \times 10^{-3}\ \text{L}/\cancel{\text{mL}}$$
$$= 0.850\ \text{L}$$

Amount in moles, *n*, of NaBr(aq):

$$n_{\text{NaBr}} = c \times V$$
$$= 0.350\ \text{mol}/\cancel{\text{L}} \times 0.850\ \cancel{\text{L}}$$
$$= 0.2975\ \text{mol}$$

Amount in moles, *n*, of Br_2(g):

$$\frac{2\ \text{mol NaBr}}{1\ \text{mol Br}_2} = \frac{0.2975\ \text{mol NaBr}}{n_{\text{Br}_2}}$$

$$n_{\text{Br}_2} = \frac{1\ \text{mol Br}_2 \times 0.2975\ \cancel{\text{mol NaBr}}}{2\ \cancel{\text{mol NaBr}}}$$
$$= 0.14875\ \text{mol}$$

Molar mass, *M*, of Br_2(g):

$$M_{\text{Br}_2} = 2M_{\text{Br}}$$
$$= 2(79.90\ \text{g/mol})$$
$$= 159.90\ \text{g/mol}$$

Mass, *m*, of Br_2(g):

$$m_{\text{Br}_2} = n \times M$$
$$= 0.14875\ \cancel{\text{mol}} \times 159.90\ \text{g}/\cancel{\text{mol}}$$
$$= 23.77\ \text{g}$$
$$= 24\ \text{g}$$

The mass of the bromine gas is 24 g.

Check Your Solution

The units for amount and concentration are correct. The answer has two significant digits and seems reasonable.

17. Practice Problem (page 417)

What mass of strontium carbonate, $SrCO_3(s)$, can be precipitated from 50.0 mL of 0.165 mol/L strontium nitrate, $Sr(NO_3)_2(aq)$, by adding excess sodium carbonate, $Na_2CO_3(aq)$?

What Is Required?

You need to calculate the mass of strontium carbonate that can be precipitated in a reaction.

What Is Given?

You know the volume of the strontium nitrate solution, $Sr(NO_3)_2(aq)$: 50.0 mL
You know the initial concentration of $Sr(NO_3)_2(aq)$: 0.165 mol/L
You know the other reactant is aqueous sodium carbonate, $Na_2CO_3(aq)$.

Plan Your Strategy

Predict the other product that forms in this double displacement reaction.
Write the balanced equation for the double displacement reaction.
Convert the volume from millilitres to litres: 1 mL = 1×10^{-3} L
Calculate the amount in moles of strontium nitrate using the relationship $n = c \times V$.
Equate the mole ratios and solve for the amount in moles of the precipitate, strontium carbonate, $SrCO_3(s)$.
Use the periodic table to determine the molar mass of $Sr(CO_3)_2$.
Calculate the mass of precipitate using the relationship $m = n \times M$.

Act on Your Strategy

The other product in this reaction is sodium nitrate, $NaNO_3(aq)$.

Balanced equation:	$Sr(NO_3)_2(aq)$ +	$Na_2CO_3(aq)$ →	$2NaNO_3(aq)$ +	$SrCO_3(s)$
Mole ratio:	1 mole	1 mole	2 moles	1 mole

Volume of solution:

$$V = 50.0\ \cancel{mL} \times 1 \times 10^{-3}\ L/\cancel{mL}$$
$$= 0.050\ L$$

Amount in moles, *n*, of $Sr(NO_3)_2(aq)$:

$$n = c \times V$$
$$= 0.165\ mol/\cancel{L} \times 0.050\ \cancel{L}$$
$$= 0.00825\ mol$$

Amount in moles, *n*, of the precipitate, $SrCO_3(s)$:

$$\frac{1\ \text{mol SrCO}_3}{1\ \text{mol Sr(NO}_3)_2} = \frac{n_{\text{SrCO}_3}}{0.00825\ \text{mol Sr(NO}_3)_2}$$

$$\begin{aligned} n_{\text{SrCO}_3} &= \frac{1\ \text{mol SrCO}_3 \times 0.00825\ \cancel{\text{mol Sr(NO}_3)_2}}{1\ \cancel{\text{mol Sr(NO}_3)_2}} \\ &= 0.00825\ \text{mol} \end{aligned}$$

Molar mass, *M*, of $SrCO_3(s)$:

$$\begin{aligned} M_{\text{SrCO}_3} &= 1M_{\text{Sr}} + 1M_{\text{C}} + 3M_{\text{O}} \\ &= 1(87.62\ \text{g/mol}) + 1(12.01\ \text{g/mol}) + 3(16.00\ \text{g/mol}) \\ &= 147.63\ \text{g/mol} \end{aligned}$$

Mass, *m*, of $SrCO_3(s)$:

$$\begin{aligned} m_{\text{SrCO}_3} &= n \times M \\ &= 0.00825\ \cancel{\text{mol}} \times 147.63\ \text{g/}\cancel{\text{mol}} \\ &= 1.2179\ \text{g} \\ &= 1.22\ \text{g} \end{aligned}$$

The mass of strontium carbonate that can be precipitated is 1.22 g.

Check Your Solution

The units for amount and concentration are correct. The answer has three significant digits and seems reasonable.

18. Practice Problem (page 417)

Before it was banned in the 1970s due to its non-selective toxicity, thallium(I) sulfate, $Tl_2SO_4(s)$ was the active ingredient in some pesticides. A chemist measured 100.0 mL of a solution of thallium(I) sulfate and added excess aqueous potassium iodide to precipitate yellow thallium(I) iodide, TlI(s). The mass of the dry precipitate was 2.45 g. Find the molar concentration of the thallium(I) sulfate solution.

What Is Required?

You need to determine the initial concentration of the thallium sulfate solution.

What Is Given?

You know the volume of the thallium sulfate solution, $Tl_2SO_4(aq)$: 100 mL
You know the mass of the thallium(I) iodide precipitate: 2.45 g
You know the other reactant: potassium iodide, KI(aq)

Plan Your Strategy

Predict the name and formula for the other product in this double displacement reaction.

Write the balanced equation for the double displacement reaction.

Use the periodic table to determine the molar mass of the precipitate.

Calculate the amount in moles of TlI(s) using the relationship $n = \dfrac{m}{M}$.

Equate the mole ratios and solve for the amount in moles of Tl_2SO_4(aq).

Convert the volume from millilitres to litres: 1 mL = 1 × 10^{-3} L

Calculate the concentration of Tl_2SO_4(aq) using the relationship $c = \dfrac{n}{V}$.

Act on Your Strategy

The other product is potassium sulfate, K_2SO_4(aq).

Balanced equation:	$Tl_2SO_4(aq)$ +	$2KI(aq) \rightarrow$	$K_2SO_4(aq)$ +	$2TlI(s)$
Mole ratio:	1 mole	2 moles	1 mole	2 moles

Molar mass, **M**, of the precipitate, TlI(s):

$$\begin{aligned} M_{\text{TlI}} &= 1M_{\text{Tl}} + 1M_{\text{I}} \\ &= 1(204.38 \text{ g/mol}) + 1(126.90 \text{ g/mol}) \\ &= 331.28 \text{ g/mol} \end{aligned}$$

Amount in moles, *n*, of TlI(s):

$$\begin{aligned} n_{\text{TlI}} &= \frac{m}{M} \\ &= \frac{2.45 \cancel{\text{ g}}}{331.28 \cancel{\text{ g}}\text{/mol}} \\ &= 7.39555 \times 10^{-3} \text{ mol} \end{aligned}$$

Amount in moles, *n*, of Tl_2SO_4(aq):

$$\frac{2 \text{ mol TlI}}{1 \text{ mol Tl}_2\text{SO}_4} = \frac{7.39555 \times 10^{-3} \text{ mol TlI}}{n_{\text{Tl}_2\text{SO}_4}}$$

$$\begin{aligned} n_{\text{Tl}_2\text{SO}_4} &= \frac{1 \text{ mol Tl}_2\text{SO}_4 \times 7.39555 \times 10^{-3} \cancel{\text{ mol TlI}}}{2 \cancel{\text{ mol TlI}}} \\ &= 3.6977^{-3} \text{ mol} \end{aligned}$$

Volume of solution:

$$V = 100\ \cancel{\text{mL}} \times 1 \times 10^{-3}\ \text{L}/\cancel{\text{mL}}$$
$$= 0.100\ \text{L}$$

Concentration of thallium sulfate solution:

$$c = \frac{n}{V}$$
$$= \frac{3.6977 \times 10^{-3}\ \text{mol}}{0.100\ \text{L}}$$
$$= 3.6977 \times 10^{-2}\ \text{mol/L}$$
$$= 3.70 \times 10^{-2}\ \text{mol/L}$$

The molar concentration of the thallium sulfate solution, Tl_2SO_4(aq), is 3.70×10^{-2} mol/L.

Check Your Solution
The units for amount and concentration are correct. The answer has three significant digits and seems reasonable.

19. Practice Problem (page 417)
A sample of a substance known to contain chloride ions was dissolved in distilled water in a 1 L
volumetric flask. Then 25.00 mL of this solution was treated with excess silver nitrate, $AgNO_3$(aq). The precipitate of silver chloride, AgCl(s), was filtered and dried. The mass of the dry precipitate was 0.765 g.
a. Calculate the concentration of chloride ions in the solution.
b. If the original substance was sodium chloride, NaCl(s), what mass of it was dissolved in the volumetric flask?

a. concentration of chloride ions
What Is Required?
You need to determine the concentration of chloride ions in the solution.

What Is Given?
You know the volume of the solution containing chloride ions, Cl^-(aq): 25.00 mL
You know the mass of the silver chloride precipitate: 0.765 g
You know the other reactant: silver nitrate, $AgNO_3$ (aq)

Plan Your Strategy

Use the periodic table to determine the molar mass of the AgCl(s) precipitate.

Calculate the amount in moles of AgCl(s) using the relationship $n = \frac{m}{M}$.

Determine the amount of chloride ions, Cl^-(aq), per mole of AgCl.
Calculate the amount in moles of chloride ions, Cl^-(aq).
Convert the volume from millilitres to litres: 1 mL = 1 × 10^{-3} L

Calculate the concentration of Cl^-(aq) using the relationship $c = \frac{n}{V}$.

Act on Your Strategy

Molar mass, *M*, of the precipitate, AgCl(s):

$$\begin{aligned} M_{\text{AgCl}} &= 1M_{\text{Ag}} + 1M_{\text{Cl}} \\ &= 1(107.87\ \text{g/mol}) + 1(35.45\ \text{g/mol}) \\ &= 43.32\ \text{g/mol} \end{aligned}$$

Amount in moles, *n*, of AgCl(s):

$$\begin{aligned} n_{\text{AgCl}} &= \frac{m}{M} \\ &= \frac{0.765\ \cancel{\text{g}}}{143.32\ \cancel{\text{g}}\text{/mol}} \\ &= 5.3377 \times 10^{-3}\ \text{mol} \end{aligned}$$

Amount in moles, *n*, of Cl^-(aq):

$$\frac{1\ \text{mol Cl}^-}{1\ \text{mol AgCl}} = \frac{n_{\text{Cl}^-}}{5.3377 \times 10^{-3}\ \text{mol AgCl}}$$

$$\begin{aligned} n_{\text{Cl}^-} &= \frac{1\ \text{mol Cl}^- \times 5.3377 \times 10^{-3}\ \cancel{\text{mol AgCl}}}{1\ \cancel{\text{mol AgCl}}} \\ &= 5.3377 \times 10^{-3}\ \text{mol} \end{aligned}$$

Volume of solution:

$$\begin{aligned} V &= 25.00\ \cancel{\text{mL}} \times 1 \times 10^{-3}\ \text{L/}\cancel{\text{mL}} \\ &= 0.02500\ \text{L} \end{aligned}$$

Concentration of chloride ions, $Cl^-(aq)$, in solution:

$$\begin{aligned} c &= \frac{n}{V} \\ &= \frac{5.3377\times10^{-3}\ \text{mol}}{0.02500\ \text{L}} \\ &= 0.213508\ \text{mol/L} \\ &= 0.214\ \text{mol/L} \end{aligned}$$

The concentration of the chloride ions, $Cl^-(aq)$, is 0.214 mol/L.

b. mass of sodium chloride

What Is Required?

You need to find the mass of sodium chloride, NaCl(s), that has this amount of chloride ions.

What Is Given?

You know the chemical formula for sodium chloride: NaCl

Plan Your Strategy

Determine the mole ratio: $Cl^-(aq)$: NaCl(s)
Use this ratio to calculate the amount in moles of NaCl(s).
Determine the molar mass of NaCl(s).
Calculate the mass of NaCl(s) per litre of solution using the relationship $m = n \times M$.

Act on Your Strategy

There is 0.213508 mol of Cl^- in 1 L of solution.

Amount in moles, *n*, of NaCl(s):

$$\frac{1\ \text{mol Cl}^-}{1\ \text{mol NaCl}} = \frac{0.213508\ \text{mol Cl}^-}{n_{\text{NaCl}}}$$

$$\begin{aligned} n_{\text{NaCl}} &= \frac{1\ \text{mol NaCl} \times 0.213508\ \cancel{\text{mol Cl}^-}}{1\ \cancel{\text{mol Cl}^-}} \\ &= 0.213508\ \text{mol} \end{aligned}$$

Molar mass, *M*, of NaCl(s):

$$\begin{aligned} M_{\text{NaCl}} &= 1M_{\text{Na}} + 1M_{\text{Cl}} \\ &= 1(22.99\ \text{g/mol}) + 1(35.45\ \text{g/mol}) \\ &= 58.44\ \text{g/mol} \end{aligned}$$

Mass, *m*, of NaCl(s):

$$m_{\text{NaCl}} = 0.213508\ \cancel{\text{mol}} \times 58.44\ \text{g}/\cancel{\text{mol}}$$
$$= 12.477\ \text{g}$$
$$= 12.5\ \text{g}$$

The mass of sodium chloride in the 1.0 L volumetric flask is 12.5 g.

Check Your Solution
The units for amount and concentration are correct. The answer has three significant digits and seems reasonable.

20. Practice Problem (page 417)
Food manufacturers sometimes add calcium acetate, $Ca(CH_3COO)_2(s)$, to sauces as a thickening agent. When analyzed, a 250 mL solution of calcium acetate was found to contain 0.200 mol of acetate ions.
a. Find the molar concentration of the calcium acetate solution.
b. What mass of calcium acetate was dissolved to make the solution?

a. molar concentration
What Is Required?
You need to find the molar concentration of a calcium acetate solution.

What Is Given?
You know the chemical formula for calcium acetate: $Ca(CH_3COO)_2(s)$
You know the amount in moles of acetate ions, $CH_3COO^-(aq)$: 0.200 mol
You know the volume of the solution: 250 mL

Plan Your Strategy
Use the mole ratio of acetate ions to calcium acetate to determine the amount in moles of calcium acetate.
Convert the volume from millilitres to litres: 1 mL = 1 × 10^{-3} L
Calculate the concentration of calcium acetate using the relationship $c = \dfrac{n}{V}$.

Act on Your Strategy
The chemical formula for calcium acetate, $Ca(CH_3COO)_2(s)$, indicates that there are two CH_3COO^- ions for one formula unit of $Ca(CH_3COO)_2(s)$.

Amount in moles, n, of $Ca(CH_3COO)_2(s)$:

$$\frac{1 \text{ mol } Ca(CH_3COO)_2}{2 \text{ mol } CH_3COO^-} = \frac{n_{Ca(CH_3COO)_2}}{0.200 \text{ mol } CH_3COO^-}$$

$$n_{Ca(CH_3COO)_2} = \frac{1 \text{ mol } Ca(CH_3COO)_2 \times 0.200 \cancel{\text{ mol } CH_3COO^-}}{2 \cancel{\text{ mol } CH_3COO^-}}$$

$$= 0.100 \text{ mol}$$

Volume of solution:

$$V = 250 \cancel{\text{ mL}} \times 1 \times 10^{-3} \text{ L/}\cancel{\text{mL}}$$

$$= 0.250 \text{ L}$$

Concentration of $Ca(CH_3COO)_2(aq)$:

$$c = \frac{n}{V}$$

$$= \frac{0.100 \text{ mol}}{0.250 \text{ L}}$$

$$= 0.40 \text{ mol/L}$$

The concentration of calcium acetate is 0.40 mol/L.

b. mass of calcium acetate

What Is Required?

You need to find the mass of calcium acetate in 250 mL of solution.

What Is Given?

You know the molar concentration: 0.40 mol/L

You know the volume of solution: 0.250 L

Plan Your Strategy

Calculate the amount in moles of calcium acetate using the relationship the relationship $n = c \times V$.

Use the periodic table to find the molar mass of $Ca(CH_3COO)_2(s)$.

Calculate the mass of calcium acetate using the relationship $m = n \times M$.

Act on Your Strategy

Amount in moles, *n*, of $Ca(CH_3COO)_2(aq)$:

$$\begin{aligned} n_{Ca(CH_3COO)_2} &= c \times V \\ &= 0.40\ \text{mol}/\cancel{\text{L}} \times 0.250\ \cancel{\text{L}} \\ &= 0.10\ \text{mol} \end{aligned}$$

Molar mass, *M*, of $Ca(CH_3COO)_2(s)$:

$$\begin{aligned} M_{Ca(CH_3COO)_2} &= 1M_{Ca} + 4M_{C} + 6M_{H} + 4M_{O} \\ &= 1(40.08\ \text{g/mol}) + 4(12.01\ \text{g/mol}) + 6(1.01\ \text{g/mol}) + 4(16.00\ \text{g/mol}) \\ &= 158.18\ \text{g/mol} \end{aligned}$$

Mass, *m*, of $Ca(CH_3COO)_2(s)$:

$$\begin{aligned} m_{Ca(CH_3COO)_2} &= n \times M \\ &= 0.10\ \cancel{\text{mol}} \times 158.18\ \text{g}/\cancel{\text{mol}} \\ &= 15.818\ \text{g} \\ &= 16\ \text{g} \end{aligned}$$

The mass of calcium acetate is 16 g.

Check Your Solution

The units for amount and concentration are correct. The answer has two significant digits and seems reasonable.

Section 9.2 Solution Stoichiometry

Solutions for Practice Problems

Student Edition page 420

21. Practice Problem (page 420)

Lead(II) sulfide, PbS(s), is a black, insoluble substance. Calculate the maximum mass of lead(II) sulfide that will precipitate when 6.75 g of sodium sulfide, $Na_2S(s)$, is added to 250 mL of 0.200 mol/L lead(II) nitrate, $Pb(NO_3)_2(aq)$.

What Is Required?

You need to find the mass of lead(II) sulfide that will precipitate.

What Is Given?
You know the mass of the sodium sulfide solution: 6.75 g
You know the volume of the lead(II) nitrate solution: 250 mL
You know the concentration of the lead(II) nitrate solution: 0.200 mol/L

Plan Your Strategy
Write the balanced chemical equation for the reaction.
Determine the molar mass of $Na_2S(s)$.
Calculate the amount in moles of each reactant.
To allow for the mole ratio of the reactants, divide the amount of each reactant by its coefficient in the chemical equation. The smaller result identifies the limiting reactant.
Use the mole ratio in the balanced equation and the amount in moles of the limiting reactant to find the amount in moles of the precipitate.
Determine the molar mass of PbS(s).
Calculate the mass of PbS(s) using the relationship $m = n \times M$.

Act on Your Strategy
Balanced chemical equation: $Na_2S(s) + Pb(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + PbS(s)$

Amount in moles, *n*, of $Na_2S(s)$:

$$\begin{aligned} n_{Na_2S} &= \frac{n}{M} \\ &= \frac{6.75\ \cancel{g}}{78.05\ \cancel{g}/mol} \\ &= 8.6483 \times 10^{-2}\ mol \end{aligned}$$

Molar mass, *M*, of $Na_2S(s)$:

$$\begin{aligned} M_{Na_2S} &= 2M_{Na} + 1M_{S} \\ &= 2(22.99\ g/mol) + 1(32.07\ g/mol) \\ &= 78.05\ g/mol \end{aligned}$$

Amount in moles, *n*, of $Pb(NO_3)_2(aq)$:

$$\begin{aligned} n_{Pb(NO_3)_2} &= c \times V \\ &= 0.200\ mol/\cancel{L} \times 0.250\ \cancel{L} \\ &= 5.00 \times 10^{-2}\ mol \end{aligned}$$

Identification of the limiting reactant:

$$\frac{\text{amount of Na}_2\text{S}}{\text{coefficient}} = \frac{8.6483 \times 10^{-2}\ \text{mol}}{1}$$
$$= 8.6483 \times 10^{-2}\ \text{mol}$$

$$\frac{\text{amount of Pb(NO}_3)_2}{\text{coefficient}} = \frac{5.00 \times 10^{-2}\ \text{mol}}{1}$$
$$= 5.00 \times 10^{-2}\ \text{mol}$$

$Pb(NO_3)_2$(aq) is the limiting reactant because it is the smaller amount.

Amount in moles, *n*, of the precipitate, PbS(s):

$$\frac{1\ \text{mol Pb(NO}_3)_2}{1\ \text{mol PbS}} = \frac{5.00 \times 10^{-2}\ \text{mol Pb(NO}_3)_2}{n_{\text{PbS}}}$$

$$n_{\text{PbS}} = \frac{1\ \text{mol PbS} \times 5.00 \times 10^{-2}\ \cancel{\text{mol Pb(NO}_3)_2}}{1\ \cancel{\text{mol Pb(NO}_3)_2}}$$
$$= 5.00 \times 10^{-2}\ \text{mol}$$

Molar mass, *M*, of PbS(s):

$$\begin{aligned} M_{\text{PbS}} &= 1M_{\text{Pb}} + 1M_{\text{S}} \\ &= 1(207.2\ \text{g/mol}) + 1(32.07\ \text{g/mol}) \\ &= 239.27\ \text{g/mol} \end{aligned}$$

Mass, *m*, of PbS(s):

$$\begin{aligned} m_{\text{PbS}} &= n \times M \\ &= 5.00 \times 10^{-2}\ \cancel{\text{mol}} \times 239.27\ \text{g/}\cancel{\text{mol}} \\ &= 11.963\ \text{g} \\ &= 12\ \text{g} \end{aligned}$$

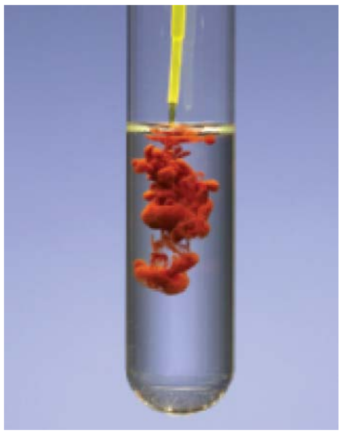
The mass of lead(II) sulfide that precipitates is 12 g.

Check Your Solution

The mass of precipitate seems reasonable compared with the number of moles of reactant used in this reaction. The answer correctly shows two significant digits.

22. Practice Problem (page 420)

Silver chromate, $Ag_2CrO_4(s)$, is a brick-red insoluble substance that is used to stain neurons so that they can be viewed under a microscope. Silver chromate can be formed by the reaction between silver nitrate, $AgNO_3(aq)$, and potassium chromate, $K_2CrO_4(aq)$, as shown in the photograph below. Calculate the mass of silver chromate that forms when 25.0 mL of 0.125 mol/L silver nitrate reacts with 20.0 mL of 0.150 mol/L potassium chromate.



What Is Required?

You need to find the mass of silver chromate that will precipitate.

What Is Given?

You know the volume of the silver nitrate solution: 25.0 mL
You know the concentration of the silver nitrate solution: 0.125 mol/L
You know the volume of the potassium chromate solution: 20.0 mL
You know the concentration of the potassium chromate solution: 0.150 mol/L

Plan Your Strategy

Write the balanced chemical equation for the reaction.
Calculate the amount in moles of each reactant using the relationship $n = c \times V$.
To allow for the mole ratio of the reactants, divide the amount of each reactant by its coefficient in the chemical equation. The smaller result identifies the limiting reactant.
Use the mole ratios to find the amount in moles, *n*, of the precipitate.
Determine the molar mass of $Ag_2CrO_4(s)$.
Calculate the mass of $Ag_2CrO_4(s)$ using the relationship $m = n \times M$.

Act on Your Strategy

Balanced chemical equation:

$2AgNO_3(aq) + K_2CrO_4(aq) \rightarrow 2KNO_3(aq) + Ag_2CrO_4(s)$

Amount in moles, *n*, of $AgNO_3(aq)$:

$$\begin{aligned} n_{AgNO_3} &= c \times V \\ &= 0.125\ \text{mol}/\cancel{\text{L}} \times 0.0250\ \cancel{\text{L}} \\ &= 3.125 \times 10^{-3}\ \text{mol} \end{aligned}$$

Amount in moles, *n*, of $K_2CrO_4(aq)$:

$$\begin{aligned} n_{K_2CrO_4} &= c \times V \\ &= 0.150\ \text{mol}/\cancel{\text{L}} \times 0.0200\ \cancel{\text{L}} \\ &= 3.000 \times 10^{-3}\ \text{mol} \end{aligned}$$

Identification of the limiting reactant:

$$\begin{aligned} \frac{\text{amount of AgNO}_3}{\text{coefficient}} &= \frac{3.125 \times 10^{-3}\ \text{mol}}{2} \\ &= 1.5625 \times 10^{-3}\ \text{mol} \end{aligned}$$

$$\begin{aligned} \frac{\text{amount of K}_2\text{CrO}_4}{\text{coefficient}} &= \frac{3.00 \times 10^{-3}}{1} \\ &= 3.000 \times 10^{-3}\ \text{mol} \end{aligned}$$

$AgNO_3(aq)$ is the limiting reactant because it is the smaller amount.

Amount in moles of the precipitate, $Ag_2CrO_4(s)$:

$$\frac{2\ \text{mol AgNO}_3}{1\ \text{mol Ag}_2\text{CrO}_4} = \frac{3.125 \times 10^{-3}\ \text{mol AgNO}_3}{n_{Ag_2CrO_4}}$$

$$\begin{aligned} n_{Ag_2CrO_4} &= \frac{\text{Ag}_2\text{CrO}_4(\text{s}) \times 3.125 \times 10^{-3}\ \cancel{\text{mol AgNO}_3}}{2\ \cancel{\text{mol AgNO}_3}} \\ &= 1.5625 \times 10^{-3}\ \text{mol} \end{aligned}$$

Molar mass, *M*, of $Ag_2CrO_4(s)$:

$$\begin{aligned} M_{Ag_2CrO_4} &= 2M_{Ag} + 1M_{Cr} + 4M_{O} \\ &= 2(107.87\ \text{g/mol}) + 1(52.00\ \text{g/mol}) + 4(16.00\ \text{g/mol}) \\ &= 331.74\ \text{g/mol} \end{aligned}$$

Mass, *m*, of $Ag_2CrO_4(s)$:

$$
\begin{aligned}
m_{Ag_2CrO_4} &= n \times M \\
&= 1.5625 \times 10^{-3}\ \cancel{mol} \times 331.74\ \ g/\cancel{mol} \\
&= 0.51834\ g \\
&= 0.518\ g
\end{aligned}
$$

The mass of silver chromate that precipitates is 0.518 g.

Check Your Solution

The mass of precipitate seems reasonable compared to the number of moles of reactant used. The answer correctly shows three significant digits.

23. Practice Problem (page 420)

Mercury compounds are poisonous, but mercury ions can be removed from a solution by precipitating insoluble mercury(II) sulfide, HgS(s). Determine the minimum volume of 0.0783 mol/L sodium sulfide, $Na_2S(aq)$, that is needed to precipitate all the mercury ions in 75.5 mL of 0.100 mol/L $Hg(NO_3)_2(aq)$.

What Is Required?

You need to calculate the volume of sodium sulfide solution needed to precipitate the mercury(II) ions from a solution of mercury(II) nitrate.

What Is Given?

You know the concentration of the sodium sulfide solution: 0.0783 mol/L
You know the volume of the mercury(II) nitrate solution: 75.5 mL
You know the concentration of the mercury(II) nitrate solution: 0.100 mol/L

Plan Your Strategy

Write the balanced chemical equation for the reaction.
Calculate the amount in moles of the reactant $Hg(NO_3)_2$ using the relationship $n = c \times V$.
Determine the amount in moles of $Hg^{2+}(aq)$.
Use the mole ratio in the balanced equation, with the determined amount of $Hg^{2+}(aq)$, to find the amount in moles of $Na_2S(aq)$.
Calculate the volume of Na_2S (aq) using the relationship $V = \dfrac{n}{c}$.

Act on Your Strategy

Balanced chemical equation:
$Na_2S(aq) + Hg(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + HgS(s)$

Amount in moles, n, of $Hg(NO_3)_2(aq)$:

$$\begin{aligned} n_{Hg(NO_3)_2} &= c \times V \\ &= 0.100\ \text{mol}/\cancel{L} \times 0.0755\ \cancel{L} \\ &= 0.007550\ \text{mol} \\ &= 7.550 \times 10^{-3}\ \text{mol} \end{aligned}$$

Amount in moles, n, of $Hg^{2+}(aq)$:

$$\frac{n_{Hg^{2+}}}{7.550 \times 10^{-3}\ \text{mol Hg(NO}_3)_2} = \frac{1\ \text{mol Hg}^{2+}}{1\ \text{mol Hg(NO}_3)_2}$$

$$\begin{aligned} n_{Hg^{2+}} &= \frac{7.550 \times 10^{-3}\ \cancel{\text{mol Hg(NO}_3)_2} \times 1\ \text{mol Hg}^{2+}}{1\ \cancel{\text{mol Hg(NO}_3)_2}} \\ &= 7.550 \times 10^{-3}\ \text{mol} \end{aligned}$$

Amount in moles, n, of $Na_2S(aq)$:

$$\frac{1\ \text{mol Hg}^{2+}}{1\ \text{mol Na}_2\text{S}} = \frac{7.550 \times 10^{-3}\ \text{mol Hg}^{2+}}{n_{Na_2S}}$$

$$\begin{aligned} n_{Na_2S} &= \frac{1\ \text{mol Na}_2\text{S} \times 7.5550 \times 10^{-3}\ \cancel{\text{mol Hg}^{2+}}}{1\ \cancel{\text{mol Hg}^{2+}}} \\ &= 7.550 \times 10^{-3}\ \text{mol} \end{aligned}$$

Volume of $Na_2S(aq)$:

$$\begin{aligned} V &= \frac{n}{c} \\ &= \frac{7.550 \times 10^{-3}\ \cancel{\text{mol}}}{0.0783\ \cancel{\text{mol}}/\text{L}} \\ &= 9.642 \times 10^{-2}\ \text{L} \\ &= 96.4\ \text{mL} \end{aligned}$$

The volume of sodium sulfide that is needed is 96.4 mL.

Check Your Solution

The two solutions react in a ratio of 1:1, and the $Na_2S(aq)$ is slightly less concentrated than the $Hg(NO_3)_2(aq)$. It is reasonable that a slightly larger volume of $Na_2S(aq)$ will be required. The answer correctly shows three significant digits.

24. Practice Problem (page 420)

What is the minimum mass of sodium carbonate, $Na_2CO_3(s)$, that is needed to precipitate all the barium ions from 50.0 mL of 0.125 mol/L barium nitrate, $Ba(NO_3)_2(aq)$?

What Is Required?

You need to calculate the minimum mass of sodium carbonate needed to precipitate barium ions from a solution of barium nitrate.

What Is Given?

You know the volume of the barium nitrate solution: 50.0 mL

You know the concentration of the barium nitrate solution: 0.125 mol/L

Plan Your Strategy

Write the balanced chemical equation for the reaction.

Calculate the amount in moles of the reactant $Ba(NO_3)_2(aq)$ using the relationship $n = c \times V$.

Determine the amount in moles of $Ba^{2+}(aq)$.

Use the mole ratio of $Ba^{2+}(aq)$ ions to $Na_2CO_3(aq)$ and the amount of $Ba^{2+}(aq)$ to find the amount in moles of $Na_2CO_3(aq)$.

Determine the molar mass of $Na_2CO_3(s)$.

Calculate the mass of $Na_2CO_3(s)$ using the relationship $m = n \times M$.

Act on Your Strategy

Balanced chemical equation:

$Na_2CO_3(aq) + Ba(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + BaCO_3(s)$

Amount in moles, *n*, of $Ba(NO_3)_2(aq)$:

$$\begin{aligned} n_{Ba(NO_3)_2} &= c \times V \\ &= 0.125\ \text{mol}/\cancel{\text{L}} \times 0.0500\ \cancel{\text{L}} \\ &= 6.250 \times 10^{-3}\ \text{mol} \end{aligned}$$

Amount in moles, *n*, of $Ba^{2+}(aq)$:

$$\frac{n_{Ba^{2+}}}{6.250 \times 10^{-3}\ \text{mol Ba}(\text{NO}_3)_2} = \frac{1\ \text{mol Ba}^{2+}}{1\ \text{mol Ba}(\text{NO}_3)_2}$$

$$\begin{aligned} n_{Ba^{2+}} &= \frac{6.250 \times 10^{-3}\ \cancel{\text{mol Ba}(\text{NO}_3)_2} \times 1\ \text{mol Ba}^{2+}}{1\ \cancel{\text{mol Ba}(\text{NO}_3)_2}} \\ &= 6.250 \times 10^{-3}\ \text{mol} \end{aligned}$$

Amount in moles, n, of $Na_2CO_3(aq)$:

$$\frac{1 \text{ mol } Ba^{2+}}{1 \text{ mol } Na_2CO_3} = \frac{6.250 \times 10^{-3} \text{ mol } Ba^{2+}}{n_{Na_2CO_3}}$$

$$\begin{aligned} n_{Na_2CO_3} &= \frac{1 \text{ mol } Na_2CO_3 \times 6.250 \times 10^{-3} \cancel{\text{mol } Ba^{2+}}}{1 \cancel{\text{mol } Ba^{2+}}} \\ &= 6.250 \times 10^{-3} \text{ mol} \end{aligned}$$

Molar mass, M, of $Na_2CO_3(s)$:

$$\begin{aligned} M_{Na_2CO_3} &= 2M_{Na} + 1M_C + 3M_O \\ &= 2(22.99 \text{ g/mol}) + 1(12.01 \text{ g/mol}) + 3(16.00 \text{ g/mol}) \\ &= 105.99 \text{ g/mol} \end{aligned}$$

Mass, m, of $Na_2CO_3(s)$:

$$\begin{aligned} m_{Na_2CO_3} &= n \times M \\ &= 6.25 \times 10^{-3} \cancel{\text{mol}} \times 105.99 \text{ g/}\cancel{\text{mol}} \\ &= 6.6254 \times 10^{-1} \text{ g} \\ &= 6.62 \times 10^{-1} \text{ g} \end{aligned}$$

The mass of sodium carbonate required is 6.62×10^{-1} g.

Check Your Solution

The mass is correctly expressed in grams and shows three significant digits. This answer seems reasonable based upon the mole ratio in the balanced equation and the quantity of reactant that has been given.

25. Practice Problem (page 420)

What is the maximum mass of lead(II) iodide, $PbI_2(s)$, that can precipitate when 40.0 mL of a 0.345 mol/L solution of lead(II) nitrate, $Pb(NO_3)_2(aq)$, is mixed with 85.0 mL of a 0.210 mol/L solution of potassium iodide, KI(aq)? Why might the actual mass precipitated be less?

What Is Required?

You need to the find the maximum mass of lead iodide that will precipitate when solutions of lead(II) nitrate and potassium iodide are mixed.

What Is Given?
You know the volume of the potassium iodide solution: 85.0 mL
You know the concentration of the potassium iodide solution: 0.210 mol/L
You know the volume of the lead(II) nitrate solution: 40.0 mL
You know the concentration of the lead(II) nitrate solution: 0.345 mol/L

Plan Your Strategy
Write the balanced chemical equation for the reaction.
Calculate the amount in moles of each reactant using the relationship the relationship $n = c \times V$.
To allow for the mole ratio of the reactants, divide the amount of each reactant by its coefficient in the chemical equation. The smaller result identifies the limiting reactant.
Use the mole ratio in the balanced equation and the amount in moles of the limiting reactant to find the amount in moles of the precipitate, PbI_2(s).
Determine the molar mass of PbI_2(s).
Calculate the mass of PbI_2(s) using the relationship $m = n \times M$.

Act on Your Strategy
Balanced chemical equation:
$2KI(aq) + Pb(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + PbI_2(s)$

Amount in moles, *n*, of KI(aq):

$$\begin{aligned} n_{\text{KI}} &= c \times V \\ &= 0.210\ \text{mol}/\cancel{\text{L}} \times 0.0850\ \cancel{\text{L}} \\ &= 1.785 \times 10^{-2}\ \text{mol} \end{aligned}$$

Amount in moles, *n*, of $Pb(NO_3)_2$(aq):

$$\begin{aligned} n_{\text{Pb(NO}_3)_2} &= c \times V \\ &= 0.345\ \text{mol}/\cancel{\text{L}} \times 0.0400\ \cancel{\text{L}} \\ &= 1.380 \times 10^{-2}\ \text{mol} \end{aligned}$$

Identification of limiting reactant:

$$\begin{aligned} \frac{\text{amount of KI}}{\text{coefficient}} &= \frac{1.785 \times 10^{-2}\ \text{mol}}{2} \\ &= 8.925 \times 10^{-3}\ \text{mol} \end{aligned}$$

$$\begin{aligned} \frac{\text{amount of Pb(NO}_3)_2}{\text{coefficient}} &= \frac{1.380 \times 10^{-2}\ \text{mol}}{1} \\ &= 1.380 \times 10^{-2}\ \text{mol} \end{aligned}$$

KI(aq) is the limiting reactant because it is the smaller amount.

Amount in moles, *n*, of the precipitate, $PbI_2(s)$:

$$\frac{2\ \text{mol KI}}{1\ \text{mol PbI}_2} = \frac{1.785 \times 10^{-2}\ \text{mol KI}}{n_{\text{PbI}_2}}$$

$$n_{\text{PbI}_2} = \frac{1\ \text{mol PbI}_2 \times 1.785 \times 10^{-2}\ \cancel{\text{mol KI}}}{2\ \cancel{\text{mol KI}}}$$

$$= 8.925 \times 10^{-3}\ \text{mol}$$

Molar mass, *M*, of $PbI_2(s)$:

$$M_{\text{PbI}_2} = 1M_{\text{Pb}} + 2M_{\text{I}}$$

$$= 1(207.2\ \text{g/mol}) + 2(126.90\ \text{g/mol})$$

$$= 461.00\ \text{g/mol}$$

Mass, *m*, of $PbI_2(s)$:

$$m_{\text{PbI}_2} = n \times M$$

$$= 8.925 \times 10^{-3}\ \cancel{\text{mol}} \times 461.00\ \text{g/}\cancel{\text{mol}}$$

$$= 4.1144\ \text{g}$$

$$= 4.11\ \text{g}$$

The maximum mass of lead(II) iodide that precipitates is 4.11 g.

One reason that the mass of $PbI_2(s)$ could be less than this amount is that a small amount of solid lead iodide dissolves. In addition, if the mass of precipitate was obtained by filtering, some of the precipitate could pass through the filter paper with the filtrate, resulting in a lower recovered mass of precipitate.

Check Your Solution

The mass of precipitate seems reasonable compared with the amount in moles of the reactant used in this reaction. The answer correctly shows three significant digits.

26. Practice Problem (page 420)

Carbonates react with dilute hydrochloric acid to generate carbon dioxide gas. What volume of 2.00 mol/L hydrochloric acid is needed to react with 3.35 g of calcium carbonate?

What Is Required?

You need to calculate the volume of hydrochloric acid solution needed to react with a mass of calcium carbonate.

What Is Given?

You know the mass of the calcium carbonate: 3.35 g

You know the concentration of the hydrochloric acid: 2.00 mol/L

Plan Your Strategy

Write the balanced chemical equation for the reaction.

Determine the molar mass of $CaCO_3$(s).

Calculate the amount in moles of the reactant $CaCO_3$(s) using the relationship $n = \frac{m}{M}$.

Use the mole ratio in the balanced equation and the amount in moles of $CaCO_3$(s) to find the amount in moles of HCl(aq).

Calculate the volume of HCl(aq) using the relationship the relationship $V = \frac{n}{c}$.

Act on Your Strategy

Balanced chemical equation:

$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(\ell) + CO_2(g)$

Molar mass of $CaCO_3$(s):

$$\begin{aligned} M_{\mathrm{CaCO_3}} &= 1M_{\mathrm{Ca}} + 1M_{\mathrm{C}} + 3M_{\mathrm{O}} \\ &= 1(40.08\ \text{g/mol}) + 1(12.01\ \text{g/mol}) + 3(16.00\ \text{g/mol})\ 3M_{\mathrm{O}} \\ &= 100.09\ \text{g / mol} \end{aligned}$$

Amount in moles, *n*, of $CaCO_3$(aq):

$$\begin{aligned} n_{\mathrm{CaCO_3}} &= \frac{m}{M} \\ &= \frac{3.35\ \cancel{\text{g}}}{100.09\ \cancel{\text{g}}/\text{mol}} \\ &= 3.3469 \times 10^{-2}\ \text{mol} \end{aligned}$$

Amount in moles, *n*, of HCl(aq):

$$\frac{1 \text{ mol } CaCO_3}{2 \text{ mol HCl}} = \frac{3.3469 \times 10^{-2} \text{ mol } CaCO_3}{n_{HCl}}$$

$$\begin{aligned} n_{HCl} &= \frac{2 \text{ mol HCl} \times 3.3469 \times 10^{-2} \cancel{\text{mol } CaCO_3}}{1 \cancel{\text{mol } CaCO_3}} \\ &= 6.699 \times 10^{-2} \text{ mol HCl (aq)} \end{aligned}$$

Volume of HCl(aq):

$$\begin{aligned} V &= \frac{n}{c} \\ &= \frac{6.699 \times 10^{-2} \cancel{\text{mol}}}{2.00 \cancel{\text{mol}}/\text{L}} \\ &= 3.3496 \times 10^{-2} \text{ L} \\ &= 3.35 \times 10^{-2} \text{ L} \\ &= 33.5 \text{ mL} \end{aligned}$$

The volume of hydrochloric acid that is needed is 33.5 mL.

Check Your Solution
The units are correct and cancel correctly. The answer seems reasonable for this amount of reactants that have been given. The answer correctly shows three significant digits.

27. Practice Problem (page 420)
A 15.8 g strip of zinc metal was placed in 100.0 mL of silver nitrate, $AgNO_3$(aq). When the reaction was complete, the strip of zinc had a mass of 13.1 g. What was the concentration of the silver nitrate solution?

What Is Required?
You need to calculate the concentration of a silver nitrate solution.

What Is Given?
You know the initial mass of the zinc: 15.8 g
You know the final mass of the zinc: 13.1 g
You know the volume of the silver nitrate solution: 100.0 mL

Plan Your Strategy
Write the balanced chemical equation for the reaction.

Calculate the amount in moles of the zinc metal using the relationship $n = \dfrac{m}{M}$.

Use the mole ratio in the balanced equation and the amount in moles of zinc to find the amount in moles of $AgNO_3$(aq).

Calculate the concentration of $AgNO_3$(aq) using the relationship $c = \dfrac{n}{V}$.

Act on Your Strategy
Balanced chemical equation:
$2AgNO_3(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + 2\ Ag(s)$

Mass of zinc that reacts = initial mass – final mass
= 15.8 g – 13.1 g
= 2.7 g

Amount in moles, *n*, of reacted Zn(s):

$$\begin{aligned} n_{Zn} &= \frac{m}{M} \\ &= \frac{2.7\ \cancel{g}}{65.38\ \cancel{g}/\text{mol}} \\ &= 4.1297 \times 10^{-2}\ \text{mol} \end{aligned}$$

Amount in moles, *n*, of $AgNO_3$(aq):

$$\frac{1\ \text{mol Zn}}{2\ \text{mol AgNO}_3} = \frac{4.1297 \times 10^{-2}\ \text{mol Zn}}{n_{AgNO_3}}$$

$$\begin{aligned} n_{AgNO_3} &= \frac{2\ \text{mol AgNO}_3 \times 4.1297 \times 10^{-2}\ \cancel{\text{mol Zn}}}{1\ \cancel{\text{mol Zn}}} \\ &= 8.2594 \times 10^{-2}\ \text{mol} \end{aligned}$$

Concentration of $AgNO_3$(aq):

$$\begin{aligned} c &= \frac{n}{V} \\ &= \frac{8.2594 \times 10^{-2}\ \text{mol}}{0.10000\ \text{L}} \\ &= 0.82594 \\ &= 0.826\ \text{mol/L} \end{aligned}$$

The concentration of the silver nitrate solution was 0.826 mol/L.

Check Your Solution
The units are correct and cancel correctly. The answer seems reasonable for the amount of reactants that have been given. The answer correctly shows three significant digits.

28. Practice Problem (page 420)
Vinegar is an aqueous solution of acetic acid, $CH_3COOH(aq)$. What volume of 1.07 mol/L aqueous sodium hydroxide will completely react with 25.0 mL of 0.833 mol/L household vinegar?

What Is Required?
You need to calculate the volume of sodium hydroxide solution.

What Is Given?
You know the concentration of the sodium hydroxide solution: 1.07 mol/L
You know the volume of the vinegar solution: 25.0 mL
You know the concentration of the vinegar solution: 0.833 mol/L

Plan Your Strategy
Write the balanced chemical equation for the reaction.
Calculate the amount in moles of the vinegar using the relationship $n = c \times V$.
Use the mole ratio in the balanced equation and the amount in moles of vinegar to find the amount in moles of NaOH(aq).
Calculate the volume of NaOH(aq) using the relationship $V = \dfrac{n}{c}$.

Act on Your Strategy
Balanced chemical equation:
$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(\ell)$

Amount in moles, *n*, of $CH_3COOH(aq)$:

$$\begin{aligned} n_{CH_3COOH} &= c \times V \\ &= 0.833 \text{ mol/}\cancel{\text{L}} \times 0.0250 \ \cancel{\text{L}} \\ &= 2.0825 \times 10^{-2} \text{ mol} \end{aligned}$$

Amount in moles, *n*, of NaOH(aq):

$$\frac{1\ \text{mol}\ CH_3COOH}{1\ \text{mol NaOH}} = \frac{2.0825\times10^{-2}\ \text{mol}\ CH_3COOH}{n_{\text{NaOH}}}$$

$$n_{\text{NaOH}} = \frac{1\ \text{mol NaOH} \times 2.0825 \times 10^{-2}\ \cancel{\text{mol}\ CH_3COOH}}{1\ \cancel{\text{mol}\ CH_3COOH}}$$

$$= 2.0825 \times 10^{-2}\ \text{mol}$$

Volume of NaOH(aq):

$$V = \frac{n}{c}$$

$$= \frac{2.0825 \times 10^{-2}\ \cancel{\text{mol}}}{1.07\ \cancel{\text{mol}}/\text{L}}$$

$$= 1.9462 \times 10^{-2}\ \text{L}$$

$$= 19.5\ \text{mL}$$

The volume of sodium hydroxide solution that is needed is 19.5 mL.

Check Your Solution

The two solutions react in a ratio of 1:1, and the NaOH(aq) is slightly more concentrated than the CH_3COOH(aq). It is reasonable that a slightly smaller volume of NaOH(aq) will be required. The answer correctly shows three significant digits.

29. Practice Problem (page 420)

Before toothpaste was invented, people sometimes used calcium carbonate, $CaCO_3$(s), to clean their teeth. What mass of calcium carbonate can be precipitated by reacting 80.0 mL of a 0.100 mol/L solution of sodium carbonate, Na_2CO_3(aq), with 50.0 mL of a 0.100 mol/L solution of calcium chloride, $CaCl_2$(aq)?

What Is Required?

You need to find the mass of calcium carbonate that will precipitate.

What Is Given?

You know the volume of the sodium carbonate solution: 80.0 mL
You know the concentration of the sodium carbonate solution: 0.100 mol/L
You know the volume of the calcium chloride solution: 50.0 mL
You know the concentration of the calcium chloride solution: 0.100 mol/L

Plan Your Strategy

Write the balanced chemical equation for the reaction.
Calculate the amount in moles of each reactant using the relationship the relationship $n = c \times V$.
To allow for the mole ratio of the reactants, divide the amount of each reactant by its coefficient in the chemical equation. The smaller result identifies the limiting reactant.
Use the mole ratios and the amount in moles of the limiting reactant to find the amount in moles, n_2, of the precipitate.
Determine the molar mass of $CaCO_3$(s).
Calculate the mass of $CaCO_3$(s) using the relationship the relationship $m = n \times M$.

Act on Your Strategy

Balanced chemical equation:
$Na_2CO_3(aq) + CaCl_2(aq) \rightarrow 2NaCl(aq) + CaCO_3(s)$

Amount in moles, *n*, of Na_2CO_3(aq):

$$\begin{aligned} n_{Na_2CO_3} &= c \times V \\ &= 0.100 \text{ mol}/\cancel{L} \times 0.0800 \ \cancel{L} \\ &= 8.000 \times 10^{-3} \text{ mol} \end{aligned}$$

Amount in moles, *n*, of $CaCl_2$(aq):

$$\begin{aligned} n_{CaCl_2} &= c \times V \\ &= 0.100 \text{ mol}/\cancel{L} \times 0.0500 \ \cancel{L} \\ &= 5.000 \times 10^{-3} \text{ mol} \end{aligned}$$

Identification of the limiting reactant:

$$\begin{aligned} \frac{\text{amount of } Na_2CO_3}{\text{coefficient}} &= \frac{8.000 \times 10^{-3} \text{ mol}}{1} \\ &= 8.000 \times 10^{-3} \text{ mol} \end{aligned}$$

$$\begin{aligned} \frac{\text{amount of } CaCl_2}{\text{coefficient}} &= \frac{5.00 \times 10^{-3}}{1} \\ &= 5.000 \times 10^{-3} \text{ mol} \end{aligned}$$

$CaCl_2$(aq) is the limiting reactant because it is the smaller amount.

Amount in moles, *n*, of the precipitate, $CaCO_3(s)$:

$$\frac{1 \text{ mol CaCl}_2}{1 \text{ mol CaCO}_3} = \frac{5.000 \times 10^{-3} \text{ mol CaCl}_2}{n_{\text{CaCO}_3}}$$

$$n_{\text{CaCO}_3} = \frac{1 \text{ mol CaCO}_3 \times 5.000 \times 10^{-3} \cancel{\text{mol CaCl}_2}}{1 \cancel{\text{mol CaCl}_2}}$$

$$= 5.000 \times 10^{-3} \text{ mol}$$

Molar mass, *M*, of $CaCO_3(s)$:

$$M_{\text{CaCO}_3} = 1M_{\text{Ca}} + 1M_{\text{C}} + 3M_{\text{O}}$$

$$= 1(40.08 \text{ g/mol}) + 1(12.01 \text{ g/ mol}) + 3(16.00 \text{ g/mol})$$

$$= 100.09 \text{ g / mol}$$

Mass, *m*, of $CaCO_3(s)$:

$$m_{\text{CaCO}_3} = n \times M$$

$$= 5.000 \times 10^{-3} \cancel{\text{mol}} \times 100.09 \text{ g/} \cancel{\text{mol}}$$

$$= 0.500 \text{ g}$$

The mass of calcium carbonate that precipitates is 0.500 g.

Check Your Solution

The mass of precipitate seems reasonable compared with the amount in moles of reactant used. The answer correctly shows three significant digits.

30. Practice Problem (page 420)

Barium chromate, $BaCrO_4(s)$, is an insoluble yellow solid. Determine the concentration of barium ions in a solution made by mixing 50.0 mL of a 0.150 mol/L solution of barium nitrate, $Ba(NO_3)_2(aq)$, with 50.0 mL of a 0.120 mol/L solution of potassium chromate, $K_2CrO_4(aq)$.

What Is Required?

You need to find the concentration of barium ions remaining in a solution.

What Is Given?

You know the volume of the barium nitrate solution: 50.0 mL
You know the concentration of the barium nitrate solution: 0.150 mol/L
You know the volume of the potassium chromate solution: 50.0 mL
You know the concentration of the potassium chromate solution: 0.120 mol/L

Plan Your Strategy
Write the balanced chemical equation for the reaction.
Calculate the amount in moles of each reactant using the relationship $n = c \times V$.
To allow for the mole ratio of the reactants, divide the amount of each reactant by its coefficient in the chemical equation. The smaller result identifies the limiting reactant.
Determine the amount in moles of $Ba(NO_3)_2$(aq) in excess.
Determine the amount in moles of Ba^{2+}(aq) per mole of $Ba(NO_3)_2$(aq).
Determine the amount in moles of Ba^{2+} in excess.
Calculate the total volume of the mixture and determine the concentration of Ba^{2+}(aq).

Act on Your Strategy

Balanced equation:	$Ba(NO_3)_2(aq)$ +	$K_2CrO_4(aq)$ →	$2KNO_3(aq)$ +	$BaCrO_4(s)$
Mole ratio	1 mole	1 mole	2 mole	1 mole

Amount in moles, *n*, of $Ba(NO_3)_2$(aq):

$$\begin{aligned} n_{Ba(NO_3)_2} &= c \times V \\ &= 0.150 \text{ mol}/\cancel{L} \times 0.0500\ \cancel{L} \\ &= 7.50 \times 10^{-3} \text{mol} \end{aligned}$$

Amount in moles, *n*, of K_2CrO_4(aq):

$$\begin{aligned} n_{K_2CrO_4} &= c \times V \\ &= 0.120 \text{ mol}/\cancel{L} \times 0.0500\ \cancel{L} \\ &= 6.00 \times 10^{-3} \text{mol} \end{aligned}$$

$$= 0.120 \text{ mol}/\cancel{L} \times 0.0500\ \cancel{L}$$

Identification of the limiting reactant:

$$\begin{aligned} \frac{\text{amount of } Ba(NO_3)_2}{\text{coefficient}} &= \frac{7.50 \times 10^{-3} \text{ mol}}{1} \\ &= 7.50 \times 10^{-3} \text{ mol} \end{aligned}$$

$$\begin{aligned} \frac{\text{amount of } K_2CrO_4}{\text{coefficient}} &= \frac{6.00 \times 10^{-3} \text{ mol}}{1} \\ &= 6.00 \times 10^{-3} \text{ mol} \end{aligned}$$

$Ba(NO_3)_2$(aq) is the limiting reactant because it is the smaller amount.

Since the mole ratio of $Ba(NO_3)_2(aq)$ to $K_2CrO_4(aq)$ is 1:1, the excess amount in moles, *n*, of $Ba(NO_3)_2(aq)$ is the difference between the amount in moles of the two reactants.

$$
\begin{aligned}
n_{\text{excess Ba(NO}_3)_2} &= n_{\text{Ba(NO}_3)_2} - n_{\text{K}_2\text{CrO}_4} \\
&= \left(7.50 \times 10^{-3}\,\text{mol}\right) - \left(6.00 \times 10^{-3}\,\text{mol}\right) \\
&= 1.5 \times 10^{-3}\,\text{mol}
\end{aligned}
$$

Amount in moles, *n*, of $Ba^{2+}(aq)$:

$$\frac{n_{\text{Ba}^{2+}}}{1.50 \times 10^{-3}\,\text{mol Ba}\left(\text{NO}_3\right)_2} = \frac{1\ \text{mol Ba}^{2+}}{1\ \text{mol Ba}\left(\text{NO}_3\right)_2}$$

$$
\begin{aligned}
n_{\text{Ba}^{2+}} &= \frac{1.50 \times 10^{-3}\ \cancel{\text{mol Ba}\left(\text{NO}_3\right)_2} \times 1\ \text{mol Ba}^{2+}}{1\ \cancel{\text{mol Ba}\left(\text{NO}_3\right)_2}} \\
&= 1.50 \times 10^{-3}\ \text{mol}
\end{aligned}
$$

Total volume of mixture:

$$
\begin{aligned}
V &= 50.0\ \text{mL} + 50.0\ \text{mL} \\
&= 100.0\ \text{mL} \\
&= 0.100\ \text{L}
\end{aligned}
$$

Concentration of $Ba^{2+}(aq)$:

$$
\begin{aligned}
c &= \frac{n}{V} \\
&= \frac{1.50 \times 10^{-3}\ \text{mol}}{0.100\ \text{L}} \\
&= 1.50 \times 10^{-2}\ \text{mol/L}
\end{aligned}
$$

The concentration of barium ions remaining in solution is 1.50×10^{-2} mol/L.

Check Your Solution

The units in the calculations have cancelled properly and the final unit is correct. The concentration of the barium ion in excess seems reasonable. The answer correctly shows three significant digits.

Section 9.2 Solution Stoichiometry

Solutions for Selected Review Questions

Student Edition page 421

1. Review Question (page 421)

Which solution has the greater concentration of chloride ions: 0.10 mol/L magnesium chloride, $MgCl_2(aq)$, or 0.15 mol/L sodium chloride, NaCl(aq)? Explain your reasoning.

Magnesium chloride and sodium chloride are both soluble compounds that will dissociate into ions in aqueous solution.

Balanced equation:	$MgCl_2(s) \rightarrow$	$Mg^{2+}(aq)$ +	$2Cl^-(aq)$
Mole ratio:	0.10 mole	0.10 mole	0.20 mole
Balanced equation:	$NaCl(s) \rightarrow$	$Na^+(aq)$ +	$Cl^-(aq)$
Mole ratio:	0.15 mole	0.15 mole	0.15 mole

By comparing the mole ratios of $Cl^-(aq)$, the solution of 0.10 mol/L $MgCl_2(aq)$ has a greater concentration of $Cl^-(aq)$.

2. Review Question (page 421)

Calculate the molar concentration of iodide ions in each aqueous solution.

a. 15.0 g of potassium iodide dissolved in 200 mL of solution

b. 12.0 g of calcium iodide dissolved in 180 mL of solution

What Is Required?

You need to find the molar concentration, *c*, of the iodide ions, $I^-(aq)$, in each solution.

What Is Given?

You know the mass of each compound and the volume of each solution.

a. 15.0 g of potassium iodide, KI(s); 200 mL KI(aq)

b. 12.0 g of calcium iodide, $CaI_2(s)$; 180 mL $CaI_2(aq)$

Plan Your Strategy

Use the periodic table to determine the molar mass of each compound.

Calculate the amount in moles of each compound using the relationship

$n = \frac{m}{M}$.

Convert the volume from millilitres to litres: 1 mL = 1×10^{-3} L

Calculate the concentration of each compound using the relationship $c = \frac{n}{V}$.

Write the balanced chemical equation for the dissolution of each compound.

Equate the mole ratios and cross multiply to solve for *c*, the concentration of $I^-(aq)$.

Act on Your Strategy

a. potassium iodide

Molar mass, *M*, of KI(s):

$$M_{KI} = 1M_K + 1M_I$$
$$= 1(39.10 \text{g/mol}) + 1(126.90 \text{ g/mol})$$
$$= 166.0 \text{g / mol}$$

Amount in moles, *n*, of KI(s):

$$n_{KI} = \frac{n}{M}$$
$$= \frac{15.0 \cancel{\text{g}}}{166.0 \cancel{\text{g}}\text{/mol}}$$
$$= 9.036 \times 10^{-2} \text{mol}$$

Balanced equation: $KI(s) \rightarrow K^+(aq) + I^-(aq)$

Mole ratio: 1 mole 1 mole 1 mole

Amount in moles, *n*, of $I^-(aq)$:

$$\frac{1 \text{ mol KI}}{1 \text{ mol } I^-} = \frac{9.036 \times 10^{-2} \text{ mol KI}}{n_{I^-}}$$
$$n_{I^-} = \frac{1 \text{ mol } I^- \times 9.036 \times 10^{-2} \cancel{\text{mol KI}}}{1 \cancel{\text{mol KI}}}$$
$$= 9.036 \times 10^{-2} \text{ mol}$$

Volume of KI(aq):

$$V = 200 \cancel{\text{mL}} \times 1 \times 10^{-3} \text{ L/}\cancel{\text{mL}}$$
$$= 0.200 \text{ L}$$

Concentration of $I^-(aq)$:

$$c = \frac{n}{V}$$
$$= \frac{9.036 \times 10^{-2} \text{ mol}}{0.200 \text{ L}}$$
$$= 0.4518 \text{ mol/L}$$
$$= 0.5 \text{ mol/L}$$

The concentration of iodide ions, $I^-(aq)$, in the potassium iodide solution is 0.5 mol/L.

b. calcium iodide
Molar mass, M, of $CaI_2(s)$:

$$\begin{aligned}M_{CaI_2} &= 1M_{Ca} + 2M_{I}\\ &= 1(40.08\text{ g/mol}) + 2(126.90\text{ g/mol})\\ &= 293.88\text{ g/mol}\end{aligned}$$

Amount in moles, n, of $CaI_2(s)$:

$$\begin{aligned}n_{CaI_2} &= \frac{m}{V}\\ &= \frac{12.0\ \cancel{\text{g}}}{293.88\ \cancel{\text{g}}\text{/mol}}\\ &= 4.0832 \times 10^{-2}\text{ mol}\end{aligned}$$

Balanced equation: $CaI_2(s) \rightarrow Ca^{2+}(aq) + 2I^{-}(aq)$
Mole ratio: 1 mole 1 mole 2 moles

Amount in moles, n, of $I^{-}(aq)$:

$$\frac{1\text{ mol CaI}_2}{2\text{ mol I}^-} = \frac{4.0832 \times 10^{-2}\text{ mol KI}}{n_{I^-}}$$

$$\begin{aligned}n_{I^-} &= \frac{2\text{ mol I}^- \times 4.0832 \times 10^{-2}\ \cancel{\text{mol CaI}_2}}{1\ \cancel{\text{mol CaI}_2}}\\ &= 8.166 \times 10^{-2}\text{ mol}\end{aligned}$$

Volume of $CaI_2(aq)$:

$$\begin{aligned}V &= 180\ \cancel{\text{mL}} \times 1 \times 10^{-3}\text{ L/}\cancel{\text{mL}}\\ &= 0.180\text{ L}\end{aligned}$$

Concentration of $I^{-}(aq)$:

$$\begin{aligned}c &= \frac{n}{V}\\ &= \frac{8.166 \times 10^{-2}\text{ mol}}{0.180\text{ L}}\\ &= 0.4537\text{ mol/L}\\ &= 0.45\text{ mol/L}\end{aligned}$$

The concentration of iodide ions, $I^{-}(aq)$, in the calcium iodide solution is 0.45 mol/L.

Check Your Solution
The units for amount and concentration are correct. The answers have the correct number of significant digits to agree with the given data and the answers seem reasonable.

3. Review Question (page 421)
What is the minimum volume of 0.220 mol/L calcium chloride, $CaCl_2(aq)$, that is needed to precipitate all the silver ions in 110 mL of 0.166 mol/L silver nitrate, $AgNO_3(aq)$?

What Is Required?
You need to find the minimum volume of calcium chloride, $CaCl_2(aq)$, to precipitate all the silver ions in silver nitrate, $AgNO_3(aq)$.

What Is Given?
You know the volume and concentration of the silver nitrate:
110 mL of 0.166 mol/L $AgNO_3(aq)$
You know the concentration of the calcium chloride solution:
0.220 mol/L $CaCl_2(aq)$

Plan Your Strategy
Write the balanced chemical equation for the double displacement reaction.
Convert the volume of $AgNO_3(aq)$ from millilitres to litres.
Calculate the amount in moles of $AgNO_3(aq)$ using the relationship $n = c \times V$.
Equate the mole ratios and cross multiply to solve for *n*, the amount in moles of $CaCl_2(aq)$ that will react completely with the $AgNO_3(aq)$.
Calculate the volume of $CaCl_2(aq)$ that will contain this amount in moles using the relationship the relationship $V = \dfrac{n}{c}$.

Act on Your Strategy
Balanced chemical equation:
$CaCl_2(aq) + 2AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2AgCl(s)$

Volume of $AgNO_3(aq)$:
$V = 110\ \cancel{\text{mL}} \times 1 \times 10^{-3}\ \text{L}/\cancel{\text{mL}}$
$= 0.110\ \text{L}$

Amount in moles, *n*, of $AgNO_3$(aq):

$$
\begin{aligned}
n_{AgNO_3} &= c \times V \\
&= 0.166 \text{ mol}/\cancel{\text{L}} \times 0.110 \ \cancel{\text{L}} \\
&= 1.826 \times 10^{-2} \text{ mol}
\end{aligned}
$$

Amount in moles, *n*, of $CaCl_2$(aq):

$$
\frac{1 \text{ mol } CaCl_2}{2 \text{ mol } AgNO_3} = \frac{n_{CaCl_2}}{1.826 \times 10^{-2} \text{ mol } AgNO_3}
$$

$$
\begin{aligned}
n_{CaCl_2} &= \frac{1 \text{ mol } CaCl_2 \times 1.826 \times 10^{-2} \ \cancel{\text{mol } AgNO_3}}{2 \ \cancel{\text{mol } AgNO_3}} \\
&= 9.130 \times 10^{-3} \text{ mol}
\end{aligned}
$$

Volume of $CaCl_2$(aq):

$$
\begin{aligned}
V &= \frac{n}{c} \\
&= \frac{9.13 \times 10^{-3} \ \cancel{\text{mol}}}{0.220 \ \cancel{\text{mol}}/\text{L}} \\
&= 0.0415 \text{ L} \\
&= 42 \text{ mL}
\end{aligned}
$$

The minimum volume of calcium chloride solution is 42 mL.

Check Your Solution

The units for amount and concentration are correct. The answer has two significant digits and seems reasonable.

4. Review Question (page 421)

Lead(II) acetate is a poisonous compound. It is used as a colour additive in hair dyes. What volume of 1.25 mol/L lead(II) acetate, $Pb(CH_3COO)_2$(aq), contains 0.500 mol of lead(II) ions, Pb^{2+}(aq)?

Balanced equation:	$Pb(CH_3COO)_2$(aq) →	Pb^{2+}(aq) +	$2CH_3COO^-$(aq)
Mole ratio:	1 mole	1 mole	2 moles
Concentration:	1.25 mol/L	1.25 mol/L	2.50 mol/L

Volume of solution that contains 0.500 mol of Pb^{2+}(aq):

$$V = \frac{n}{c}$$
$$= \frac{0.500\ \cancel{\text{mol}}}{1.25\ \cancel{\text{mol}}/\text{L}}$$
$$= 0.400\ \text{L}$$

The volume of lead(II) acetate needed is 0.400 L.

5. Review Question (page 421)

A piece of iron was added to a beaker that contained 0.585 mol/L copper(II) sulfate, $CuSO_4$ (aq). The solid copper that precipitated was dried, and its mass was found to be 5.02 g. Some unreacted iron remained in the beaker. Calculate the minimum volume of the copper(II) sulfate solution.

What Is Required?

You need to find the minimum volume of copper(II) sulfate solution required to produce a given mass of copper.

What Is Given?

You know the concentration of the copper(II) sulfate solution: 0.585 mol/L
You know the mass of the copper precipitated: 5.02 g

Plan Your Strategy

Write the balanced chemical equation for the single displacement reaction.
Use the periodic table to determine the molar mass of Cu(s).
Calculate the amount in moles of Cu(s) using the relationship $n = \frac{m}{M}$.
Calculate the amount in moles of copper(II) sulfate using the relationship the relationship the mole ratio in the balanced equation.
Calculate the volume of copper(II) sulfate using the relationship the relationship $V = \frac{n}{c}$.

Act on Your Strategy

Balanced equation:	$Fe(s) + CuSO_4(aq) \rightarrow$	$FeSO_4(aq) +$	$Cu(s)$
Mole ratio:	1 mole		1 mole

Molar mass, M, of Cu(s)

$M_{\text{Cu}} = 63.55$ g/mol (from the periodic table)

Amount in moles, *n*, of Cu(s):

$$n_{\text{Cu}} = \frac{m}{M}$$
$$= \frac{5.02\ \cancel{\text{g}}}{63.55\ \cancel{\text{g}}/\text{mol}}$$
$$= 7.8992 \times 10^{-2}\ \text{mol}$$

Amount in moles, *n*, of $CuSO_4$(aq):

$$\frac{1\ \text{mol Cu}}{1\ \text{mol CuSO}_4} = \frac{7.8992 \times 10^{-2}\ \text{mol Cu}}{n_{\text{CuSO}_4}}$$
$$n_{\text{CuSO}_4} = \frac{1\ \text{mol CuSO}_4 \times 7.8992 \times 10^{-2}\ \cancel{\text{mol Cu}}}{1\ \cancel{\text{mol Cu}}}$$
$$= 7.8992 \times 10^{-2}\ \text{mol}$$

Volume of copper(II) sulfate solution:

$$V = \frac{n}{c}$$
$$= \frac{7.8992 \times 10^{-2}\ \cancel{\text{mol}}}{0.585\ \cancel{\text{mol}}/\text{L}}$$
$$= 0.135\ \text{L}$$

The minimum volume of copper(II) sulfate solution needed is 0.135 L.

Check Your Solution
The units for amount and concentration are correct. The answer has three significant digits and seems reasonable.

6. Review Question (page 421)
To generate hydrogen gas, a teacher added 25.0 g of mossy zinc to 220 mL of 3.00 mol/L hydrochloric acid in an Erlenmeyer flask.
a. What mass of hydrogen gas was generated?
b. After the reaction, what was the concentration of zinc chloride, $ZnCl_2$(aq), in the flask?

What Is Required?
a. You need to calculate the mass of hydrogen gas produced in a reaction.
b. You need to calculate the concentration of a zinc chloride solution.

What Is Given?
You know the mass of the mossy zinc: 25.0 g.
You know the volume and concentration of the hydrochloric acid:
220 mL of 3.00 mol/L HCl(aq).

a. mass of hydrogen gas
Plan Your Strategy
Write the balanced equation for the single displacement reaction.
Convert the volume of HCl(aq) from millilitres to litres.
Calculate the amount in moles of Zn(s) and the amount in moles of HCl(aq).
To allow for the mole ratio of the reactants, divide the amount of each reactant by its coefficient in the chemical equation. The smaller result identifies the limiting reactant.
Calculate the mass of hydrogen using the relationship $m = n \times M$.

Act on Your Strategy
Balanced chemical equation:
$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Amount in moles, *n*, of Zn(s):

$$\begin{aligned} n_{\mathrm{Zn}} &= \frac{m}{M} \\ &= \frac{25.0\ \cancel{\mathrm{g}}}{65.38\ \cancel{\mathrm{g}}/\mathrm{mol}} \\ &= 0.3824\ \mathrm{mol} \end{aligned}$$

Volume of HCl(aq):

$$\begin{aligned} V &= 220\ \cancel{\mathrm{mL}} \times 1 \times 10^{-3}\ \mathrm{L}/\cancel{\mathrm{mL}} \\ &= 0.220\ \mathrm{L} \end{aligned}$$

Amount in moles, *n*, of HCl(aq):

$$\begin{aligned} n_{\mathrm{HCl}} &= c \times V \\ &= 3.00\ \mathrm{mol}/\cancel{\mathrm{L}} \times 0.220\ \cancel{\mathrm{L}} \\ &= 0.660\ \mathrm{mol} \end{aligned}$$

Identification of limiting reactant:

$$\begin{aligned} \frac{\text{amount of Zn}}{\text{coefficient}} &= \frac{0.3824\ \mathrm{mol}}{1} \\ &= 0.3824\ \mathrm{mol} \end{aligned}$$

$$\frac{\text{amount of HCl}}{\text{coefficient}} = \frac{0.660\text{ mol}}{2}$$
$$= 0.330\text{ mol}$$

HCl(aq) is the limiting reactant because it is the smaller amount.

Amount in moles, *n*, of H_2(g):

$$\frac{2\text{ mol HCl}}{1\text{ mol H}_2} = \frac{0.660\text{ mol HCl}}{n_{H_2}}$$
$$n_{H_2} = \frac{1\text{ mol H}_2 \times 0.660\ \cancel{\text{mol HCl}}}{2\ \cancel{\text{mol HCl}}}$$
$$= 0.330\text{ mol}$$

Mass, *m*, of H_2(g):

$$m_{H_2} = n \times M$$
$$= 0.330\ \cancel{\text{mol}} \times 2.02\text{ g}/\cancel{\text{mol}}$$
$$= 0.67\text{ g}$$

The mass of H_2(g) produced is 0.67 g.

b. concentration of the zinc chloride solution

Plan Your Strategy

Use the mole ratio in the balanced equation and the amount in moles of limiting reactant to calculate the amount in moles of $ZnCl_2$(aq) that is produced.

Calculate the concentration of $ZnCl_2$(aq).

Act on Your Strategy

Amount in moles, *n*, of $ZnCl_2$(aq):

$$\frac{1\text{ mol ZnCl}_2}{2\text{ mol HCl}} = \frac{n_{ZnCl_2}}{0.660\text{ mol HCl}}$$
$$n_{ZnCl_2} = \frac{1\text{ mol ZnCl}_2 \times 0.660\ \cancel{\text{mol HCl}}}{2\ \cancel{\text{mol HCl}}}$$
$$= 0.330\text{ mol}$$

Concentration of $ZnCl_2$(aq):

$$\begin{aligned} c &= \frac{n}{V} \\ &= \frac{0.330 \text{ mol}}{0.220 \text{ L}} \\ &= 1.5 \text{ mol/L} \end{aligned}$$

The concentration of the zinc chloride in solution is 1.5 mol/L.

Check Your Solution

The limiting reactant was determined correctly and the units for amount and concentration are correct. The answers have two significant digits and seem reasonable.

7. Review Question (page 421)

A type of stomach medication is a tablet that contains a mixture of 1.00 g of sodium hydrogen carbonate, $NaHCO_3$(s), and 1.00 g of citric acid, $H_3C_6H_5O_7$(s). When dropped into water, the chemicals in the tablet react to produce carbon dioxide gas:

$3NaHCO_3$(aq) + $H_3C_6H_5O_7$(aq) → $3CO_2$(g) + $3H_2O$(ℓ) + $Na_3C_6H_5O_7$(aq)

a. Which substance is in excess?
b. What mass of this substance remains unreacted when the tablet is dropped into a glass of water?

What Is Required?

a. You need to determine which substance is in excess.
b. You need to determine the mass of the unreacted substance that was in excess.

What Is Given?

You know the mass of the reactants: 1.00 g of sodium hydrogen carbonate, $NaHCO_3$(s), and 1.00 g of citric acid, $H_3C_6H_5O_7$(s)
You know the balanced equation for the reaction between $NaHCO_3$(aq) and $H_3C_6H_5O_7$(aq).

a. substance in excess

Plan Your Strategy

Calculate the molar masses of $NaHCO_3$(s) and $H_3C_6H_5O_7$(s).

Calculate the amount in moles of $NaHCO_3$(s) and the amount in moles of $H_3C_6H_5O_7$(s).

To allow for the mole ratio of the reactants, divide the amount of each reactant by its coefficient in the chemical equation. The larger result identifies the reactant in excess.

Act on Your Strategy

Molar mass, M, of $NaHCO_3$(s):

$$\begin{aligned} M_{\mathrm{NaHCO_3}} &= 1M_{\mathrm{Na}} + 1M_{\mathrm{H}} + 1M_{\mathrm{C}} + 3M_{\mathrm{O}} \\ &= 1(22.99\ \text{g/mol}) + 1(1.01\ \text{g/mol}) + 1(12.01\ \text{g/mol}) + 3(16.00\ \text{g/mol}) \\ &= 84.10\ \text{g/mol} \end{aligned}$$

Molar mass, M, of $H_3C_6H_5O_7$(s):

$$\begin{aligned} M_{\mathrm{H_3C_6H_5O_7}} &= 8M_{\mathrm{H}} + 6M_{\mathrm{C}} + 7M_{\mathrm{O}} \\ &= 8(1.01\ \text{g/mol}) + 6(12.01\ \text{g/mol}) + 7(16.00\ \text{g/mol}) \ = 8M_{\mathrm{H}} + 6M_{\mathrm{C}} + \\ &= 192.14\ \text{g / mol} \end{aligned}$$

Amount in moles, n, of $NaHCO_3$(s):

$$\begin{aligned} n_{\mathrm{NaHCO_3}} &= \frac{m}{M} \\ &= \frac{1.0\ \cancel{\text{g}}}{84.10\ \cancel{\text{g}}/\text{mol}} \\ &= 1.189 \times 10^{-2}\ \text{mol} \end{aligned}$$

Amount in moles, n, of $H_3C_6H_5O_7$(s):

$$\begin{aligned} n_{\mathrm{H_3C_6H_5O_7}} &= \frac{m}{M} \\ &= \frac{1.0\ \cancel{\text{g}}}{192.14\ \cancel{\text{g}}/\text{mol}} \\ &= 5.2045 \times 10^{-3}\ \text{mol} \end{aligned}$$

Identification of excess reactant:

$$\frac{\text{amount of NaHCO}_3}{\text{coefficient}} = \frac{1.189 \times 10^{-2}\ \text{mol}}{3}$$
$$= 3.9633 \times 10^{-3}\ \text{mol}$$

$$\frac{\text{amount of H}_3\text{C}_6\text{H}_5\text{O}_7}{\text{coefficient}} = \frac{5.2045 \times 10^{-3}\ \text{mol}}{1}$$
$$= 5.2045 \times 10^{-3}\ \text{mol}$$

The citric acid is in excess because it is the larger amount.

b. mass of unreacted substance in excess

Plan Your Strategy

Calculate the amount in moles of $H_3C_6H_5O_7(s)$ in excess by subtracting the amount in moles needed to react with 1.189×10^{-2} mol $NaHCO_3(s)$ from the total moles of $H_3C_6H_5O_7$ given.

Calculate the mass of excess $H_3C_6H_5O_7(s)$ using the relationship $m = n \times M$.

Act on Your Strategy

Amount in moles, n, of $H_3C_6H_5O_7(s)$ in excess:

$$n_{\text{H}_3\text{C}_6\text{H}_5\text{O}_7\ \text{in excess}} = \left(5.204 \times 10^{-3}\text{mol}\right) - \left(3.963 \times 10^{-3}\text{mol}\right)$$
$$= 1.241 \times 10^{-3}\text{mol}$$

Mass, m, of $H_3C_6H_5O_7(s)$ in excess:

$$m_{\text{H}_3\text{C}_6\text{H}_5\text{O}_7\ \text{in excess}} = n \times M$$
$$= 1.241 \times 10^{-3}\ \cancel{\text{mol}} \times 192.14\ \text{g}/\cancel{\text{mol}}$$
$$= 0.23844\ \text{g}$$
$$= 0.238\ \text{g}$$

The mass of $H_3C_6H_5O_7$ in excess is 0.238 g.

Check Your Solution

The excess reactant was determined correctly and the units for amount and concentration are correct. The answer to part b has three significant digits and seems reasonable.

8. Review Question (page 421)

When 50 mL of 0.20 mol/L sodium sulfate, $Na_2SO_4(aq)$, was mixed with 80 mL of 0.10 mol/L lead(II) acetate, $Pb(CH_3COO)_2(aq)$, a white precipitate formed. Identify the precipitate, and calculate the maximum mass of dry solid that can be collected.

What Is Required?

You need to identify and determine the mass of a precipitate.

What Is Given?

You know the volume and concentration of the reactants:
50 mL of 0.20 mol/L sodium sulfate and 80 mL of 0.10 mol/L lead(II) acetate

Plan Your Strategy

Write the balanced equation for the double displacement reaction that occurs when the reactants are mixed.
Refer to the solubility guidelines on page 363 to identify the precipitate.
Convert the volumes of each reactant from millilitres to litres.
Calculate the amount in moles of each reactant using the relationship $n = c \times V$.
To allow for the mole ratio of the reactants, divide the amount of each reactant by its coefficient in the chemical equation. The smaller result identifies the limiting reactant.
Calculate the amount in moles of precipitate using the relationship the limiting reagent and the mole ratio in the balanced equation.
Determine the molar mass of the precipitate.
Calculate the mass of the precipitate using the relationship the relationship $m = n \times M$.

Act on Your Strategy

Balanced chemical equation:
$Na_2SO_4(aq) + Pb(CH_3COO)_2(aq) \rightarrow 2NaCH_3COO(aq) + PbSO_4(s)$

The precipitate is lead(II) sulfate, $PbSO_4(s)$.

Volume of $Na_2SO_4(aq)$:
$V = 50\ \cancel{mL} \times 1 \times 10^{-3}\ L/\cancel{mL}$
$= 0.050$ L

Volume of $Pb(CH_3COO)_2(aq)$:
$V = 80\ \cancel{mL} \times 1 \times 10^{-3}\ L/\cancel{mL}$
$= 0.080$ L

Amount in moles, *n*, of $Na_2SO_4(aq)$:

$$n_{Na_2SO_4} = c \times V$$
$$= 0.20\ \text{mol}/\cancel{\text{L}} \times 0.050\ \cancel{\text{L}}$$
$$= 0.010\ \text{mol}$$

Amount in moles, *n*, of $Pb(CH_3COO)_2(aq)$:

$$n_{Pb(CH_3COO)_2} = c \times V$$
$$= 0.10\ \text{mol}/\cancel{\text{L}} \times 0.080\ \cancel{\text{L}}$$
$$= 0.0080\ \text{mol}$$

Identification of limiting reactant:

$$\frac{\text{amount of Na}_2\text{SO}_4}{\text{coefficient}} = \frac{0.010\ \text{mol}}{1}$$
$$= 0.010\ \text{mol}$$

$$\frac{\text{amount of Pb(CH}_3\text{COO)}_2}{\text{coefficient}} = \frac{0.0080\ \text{mol}}{1}$$
$$= 0.0080\ \text{mol}$$

$Pb(CH_3COO)_2(aq)$ is the limiting reactant because it is the smaller amount.

Amount in moles, *n*, of $PbSO_4(s)$:

$$\frac{1\ \text{mol PbSO}_4}{1\ \text{mol Pb(CH}_3\text{CO}_2)_2} = \frac{n_{PbSO_4}}{0.0080\ \text{mol Pb(CH}_3\text{CO}_2)_2}$$
$$n_{PbSO_4} = \frac{1\ \text{mol PbSO}_4 \times 0.0080\ \cancel{\text{mol Pb(CH}_3\text{CO}_2)_2}}{1\ \cancel{\text{mol Pb(CH}_3\text{CO}_2)_2}}$$
$$= 0.0080\ \text{mol}$$

Molar mass of $PbSO_4(s)$:

$$M_{PbSO_4} = 1M_{Pb} + 1M_{s} + 4M_{O}$$
$$= 1(207.2\ \text{g/mol}) + 1(32.07\ \text{g/mol}) + 4(16.00\ \text{g/mol})$$
$$= 303.27\ \text{g / mol}$$

Mass, m, of $PbSO_4(s)$

$$m_{PbSO_4} = n \times M$$
$$= 0.0080\ \cancel{\text{mol}} \times 303.27\ \text{g}/\cancel{\text{mol}}$$
$$= 2.426\ \text{g}$$
$$= 2\ \text{g}$$

The mass of the precipitate, $PbSO_4(s)$, is 2 g.

Check Your Solution

The limiting reactant was determined correctly and the units for amount and mass are correct. The answer has one significant digit and seems reasonable**.**

9. Review Question (page 421)

To measure the concentration of copper(II) sulfate in the water discharged from an industrial plant, a chemist measured 600 mL of the water and then added excess aqueous sodium sulfide. When dried, the precipitate of copper(II) sulfide, CuS(s), had a mass of 0.125 g. Calculate the molar concentration of copper ions in the water sample.

What Is Required?

You need to determine the initial concentration of the copper(II) ions in the water.

What Is Given?

You know the volume of the water discharge: 600 mL

You know the mass of the copper(II) sulfide precipitate, CuS(s): 0.125 g

You know the other reactant is sodium sulfide, $Na_2S(aq)$.

Plan Your Strategy

Write the balanced equation for the double displacement reaction.

Use the periodic table to determine the molar mass of the precipitate.

Calculate the amount in moles of CuS(s) using the relationship $n = \frac{m}{M}$.

Equate the mole ratios and solve for the amount in moles of copper(II) sulfate, $CuSO_4(aq)$.

Convert the volume from millimetres to litres: 1 mL = 1×10^{-3} L

Calculate the concentration of $CuSO_4(aq)$ using the relationship $c = \frac{n}{V}$.

Act on Your Strategy

Balanced equation: $CuSO_4(aq) + Na_2S(aq) \rightarrow Na_2SO_4(aq) + CuS(s)$

Mole ratio: 1 mole 1 mole

Molar mass of the precipitate, CuS(s):

$M_{CuS(s)} = 1M_{Cu} + 1M_S$
$= 1(63.55 \text{ g/mol}) + 1(32.07 \text{ g/mol})$
$= 95.62 \text{ g/mol}$

Amount in moles, *n*, of CuS(s):

$$n = \frac{m}{M}$$
$$= \frac{0.125 \cancel{\text{g}}}{95.62 \cancel{\text{g}}\text{/mol}}$$
$$= 1.3072 \times 10^{-3} \text{ mol}$$

Amount in moles, *n*, of $CuSO_4(aq)$:

$$\frac{1 \text{ mol CuSO}_4}{1 \text{ mol CuS}} = \frac{n_{\text{CuSO}_4}}{1.3072 \times 10^{-3} \text{ mol CuS}}$$
$$n_{\text{CuSO}_4} = \frac{1 \text{ mol CuSO}_4 \times 1.3072 \times 10^{-3} \cancel{\text{mol CuS}}}{1 \cancel{\text{mol CuS}}}$$
$$= 1.3072 \times 10^{-3} \text{ mol}$$

Volume of solution:

$$V = 600 \cancel{\text{mL}} \times 1 \times 10^{-3} \text{ L/}\cancel{\text{mL}}$$
$$= 0.600 \text{ L}$$

Concentration of copper(II) sulfate solution:

$$c = \frac{n}{V}$$
$$= \frac{1.3072 \times 10^{-3} \text{ mol}}{0.600 \text{ L}}$$
$$= 2.17876 \times 10^{-3} \text{ mol/L}$$
$$= 2 \times 10^{-3} \text{ mol/L}$$

From the chemical formula, the concentration of the Cu^{2+}(aq) ions is equal to the concentration of $CuSO_4$(aq). The concentration of the Cu^{2+}(aq) ions is therefore 2×10^{-3} mol/L.

Check Your Solution
The units for amount and concentration are correct. The answer has one significant digit and seems reasonable.

10. Review Question (page 421)
Mixing solutions of calcium chloride, $CaCl_2$(aq), and potassium carbonate, K_2CO_3(aq), will cause calcium carbonate, $CaCO_3$(s), to precipitate. Suppose that you have the following solutions available: 0.500 mol/L $CaCl_2$(aq) and 1.00 mol/L K_2CO_3(aq). What volume of each solution should be mixed together to form 10.0 g of calcium carbonate?

What Is Required?
You need to find the volume of 0.500 mol/L $CaCl_2$(aq) and 1.00 mol/L K_2CO_3(aq) that must be mixed together to form 10.0 g of calcium carbonate.

What Is Given?
You know the formulas for and concentrations of the reactants:
0.500 mol/L $CaCl_2$(aq) and 1.00 mol/L K_2CO_3(aq)
You know the mass of the precipitate: 10.0 g $CaCO_3$(s)

Plan Your Strategy
Write the balanced equation for the overall reaction.
Calculate the molar mass of $CaCO_3$(s).
Calculate the amount in moles that is equivalent to 10.0 g of $CaCO_3$(s).
Using the mole ratio in the balanced equation, determine the number of moles of each reactant required to produce the required amount of $CaCO_3$(s).
Calculate the volume of each reactant that is required using the relationship the relationship $V = \dfrac{n}{c}$.

Act on Your Strategy

Balanced equation:	$CaCl_2$(aq) +	K_2CO_3(aq) →	2KCl(aq) +	$CaCO_3$(s)
Mole ratio:	1 mole	1 mole	2 moles	1 mole

Molar mass, M, of $CaCO_3$(s):

$$\begin{aligned} M_{\text{CaCO}_3} &= 1M_{\text{Ca}} + 1M_{\text{C}} + 3M_{\text{O}} \\ &= 40.08 \text{ g/mol}) + 1(12.01 \text{ g/mol}) + 3(16.00 \text{ g/mol}) \\ &= 100.09 \text{ g / mol} \end{aligned}$$

Amount in moles, n, of $CaCO_3(s)$:

$$n_{CaCO_3} = \frac{m}{M}$$
$$= \frac{10.0\ \cancel{g}}{100.09\ \cancel{g}/\text{mol}}$$
$$= 9.991 \times 10^{-2}\ \text{mol}$$

Amount in moles, n, of $CaCl_2(s)$:

$$\frac{1\ \text{mol}\ CaCl_2}{1\ \text{mol}\ CaCO_3} = \frac{n_{CaCl_2}}{9.991 \times 10^{-2}\ \text{mol}\ CaCO_3}$$
$$n_{CaCl_2} = \frac{1\ \text{mol}\ CaCl_2 \times 9.991 \times 10^{-2}\ \cancel{\text{mol}\ CaCO_3}}{1\ \cancel{\text{mol}\ CaCO_3}}$$
$$= 9.991 \times 10^{-2}\ \text{mol}$$

Volume of $CaCl_2(aq)$ required:

$$V = \frac{n}{c}$$
$$= \frac{9.991 \times 10^{-2}\ \cancel{\text{mol}}}{0.500\ \cancel{\text{mol}}/\text{L}}$$
$$= 1.9982 \times 10^{-1}\ \text{L}$$
$$= 2.00 \times 10^{-1}\ \text{L}$$

Amount in mole, n, of $K_2CO_3(aq)$:

$$\frac{1\ \text{mol}\ K_2CO_3}{1\ \text{mol}\ CaCO_3} = \frac{n_{K_2CO_3}}{9.991 \times 10^{-2}\ \text{mol}\ CaCO_3}$$
$$n_{K_2CO_3} = \frac{1\ \text{mol}\ K_2CO_3 \times 9.991 \times 10^{-2}\ \cancel{\text{mol}\ CaCO_3}}{1\ \cancel{\text{mol}\ CaCO_3}}$$
$$= 9.991 \times 10^{-2}\ \text{mol}$$

Volume of K_2CO_3(aq) required:

$$V = \frac{n}{c}$$

$$= \frac{9.991 \times 10^{-2}\ \cancel{\text{mol}}}{1.00\ \cancel{\text{mol}}/\text{L}}$$

$$= 9.991 \times 10^{-2}\ \text{L}$$

$$= 9.99 \times 10^{-2}\ \text{L}$$

Check Your Solution

The units for amount and concentration are correct. The answers have three significant digits and seem reasonable since the required volumes of the reactant solutions are inversely proportional to their concentrations.

12. Review Question (page 421)

Lead poisoning can have long-lasting effects. One of the most effective treatments for lead poisoning is the ion called $EDTA^{4-}$, which stands for ethylenediaminetetraacetate. $EDTA^{4-}$ ions bond to lead(II) ions in a 1:1 ratio. A doctor determines that a child's blood has a dangerously high concentration of 1.0×10^{-5} mol/L of lead(II) ions. The doctor estimates that the child's total blood volume is about 1.6 L. Find the minimum volume of a 0.025 mol/L solution of $EDTA^{4-}$ ions that is needed to treat the child.

What Is Required?

You must find the minimum volume of $EDTA^{4-}$ needed to bond to the Pb^{2+}(aq) ions in the blood.

What Is Given?

Your know the Pb^{2+}(aq) concentration: 1.0×10^{-5} mol/L
You know the concentration of the $EDTA^{4-}$(aq): 0.025 mol/L
You know the total volume of blood in which the Pb^{2+}(aq) is found: 1.6 L
You know that Pb^{2+}(aq) and $EDTA^{4-}$(aq) react in a mole ratio of 1:1.

Plan Your Strategy

Calculate the amount in moles of Pb^{2+}(aq) using the relationship $n = c \times V$.
Using the relationship the mole ratio of 1:1, determine the amount in moles of $EDTA^{4-}$(aq).

Calculate the volume of $EDTA^{4-}$(aq) using the relationship $V = \frac{n}{c}$.

Act on Your Strategy
Amount in moles, n, of Pb^{2+}(aq):

$$n_{Pb^{2+}} = c \times V$$
$$= 1.0 \times 10^{-5}\ \text{mol}/\cancel{\text{L}} \times 1.6\ \cancel{\text{L}}$$
$$= 1.6 \times 10^{-5}\ \text{mol}$$

Because the mole ratio of bonding is 1:1, the amount in moles of $EDTA^{4-}$(aq) is also 1.6×10^{-5} mol.

Volume of $EDTA^{4-}$(aq):

$$V = \frac{n}{c}$$
$$= \frac{1.6 \times 10^{-5}\ \cancel{\text{mol}}}{0.025\ \cancel{\text{mol}}/\text{L}}$$

The volume of $EDTA^{4-}$(aq) solution is 6.4×10^{-4} L or 0.64 mL.

Check Your Solution
The units for concentration and volume are correct and the answer seems reasonable with two significant digits.

Section 9.3 Water Quality

Solution for Selected Review Question

Student Edition page 429

10. Review Question (page 429)
Use Table 9.3 to decide which of the following are unacceptable in a sample of drinking water. Which of the following should be of most concern? Explain your reasoning.
• iron ions, 0.35 mg/L
• chloride ions, 200 ppm
• benzene, 0.000 007 g/L

Ion or Compound	Concentration	MAC (mg/L)	AO (mg/L)
Fe^{2+}, Fe^{3+}	0.35 mg/L		0.3
Cl^{-}	200 ppm = 200 mg/L		250
Benzene	$0.000\ 007\ \cancel{\text{g}}/\text{L} \times 1 \times 10^{3}\ \text{mg}/\cancel{\text{g}}$ $= 0.007$ mg/L	0.005	

Benzene is of most concern since the concentration exceeds the maximum allowable concentration (MAC). The iron ion concentration is slightly greater than the aesthetic objective (AO) but poses no health threat. The chloride ion concentration is below the aesthetic objective.

Section 9.4 Water Treatment

Solution for Selected Review Question

Student Edition page 436

16. Review Question (page 436)

Lead(II) ions can be present in water that has passed through lead pipes or through pipes joined using the relationship lead-tin solder.

a. Identify a compound that could be used to precipitate lead from an aqueous solution.

b. Write the net ionic equation for the reaction.

a. identification of compound

Lead(II) ions, Pb^{2+}(aq), can be precipitated by Cl^{-}(aq), Br^{-}(aq), CO_3^{2-}(aq), or SO_4^{2-}(aq) among other anions. For example, sodium carbonate, Na_2CO_3(aq), would precipitate the lead(II) ion.

b. net ionic equation

$Pb^{2+}(aq) + CO_3^{2-}(aq) \rightarrow PbCO_3(s)$