9.3 Stoichiometry in Solution Chemistry

Section Preview/ Specific Expectations

In this section, you will

- **represent** double displacement reactions by their net ionic equations
- **write** balanced chemical equations for double displacement reactions
- **solve** stoichiometry problems that involve solutions
- **describe** and **work with** solutions that have known concentrations

Recall that stoichiometry involves calculating the amounts of reactants and products in chemical reactions. If you know the atoms or ions in a formula or a reaction, you can use stoichiometry to determine the amounts of these atoms or ions that react. Solving stoichiometry problems in solution chemistry involves the same strategies you learned in Unit 2. Calculations involving solutions sometimes require a few additional steps, however. For example, if a precipitate forms, the net ionic equation may be easier to use than the chemical equation. Also, some problems may require you to calculate the amount of a reactant, given the volume and concentration of the solution.

Take your time working through the next three Sample Problems. Make sure that you understand how to arrive at the solutions. Then try the Practice Problems on page 352.

Sample Problem

The Concentration of Ions

Problem

Calculate the concentration (in mol/L) of chloride ions in each solution.

(a) 19.8 g of potassium chloride dissolved in 100 mL of solution

(b) 26.5 g of calcium chloride dissolved in 150 mL of solution

(c) a mixture of the two solutions in parts (a) and (b), assuming that the volumes are additive

What Is Required?

(a) and **(b)** You need to find the concentration (in mol/L) of chloride ions in two different solutions.

(c) You need to find the concentration of chloride ions when the two solutions are mixed.

What Is Given?

You know that 19.8 g of potassium chloride is dissolved in 100 mL of solution. You also know that 26.5 g of calcium chloride is dissolved in 150 mL of solution.

Plan Your Strategy

(a) and **(b)** For each solution, determine the molar mass. Find the amount (in mol) using the mass and the molar mass. Write equations for the dissociation of the substance. (That is, write the total ionic equation.)

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Use the coefficients in the dissociation equation to determine the amount (in mol) of chloride ions present. Calculate the concentration (in mol/L) of chloride ions from the amount and volume of the solution.

(c) Add the amounts of chloride ions in the two solutions to find the total. Add the volumes of the solutions to find the total volume. Calculate the concentration of chloride ions (in mol/L) using the total amount (in mol) divided by the total volume (in L).

Act on Your Strategy

(a) and **(b)**

Solution	KCl	$CaCl_2$
Molar mass	$39.10 + 35.45 = 74.55$ g	$40.08 + (2 \times 35.45) = 110.98$ g
Amount (mol)	$19.8\text{ g} \times \frac{1\text{ mol}}{74.55\text{ g}} = 0.266$ mol	$26.5\text{ g} \times \frac{1\text{ mol}}{110.98\text{ g}} = 0.239$ mol
Dissociation equation	$KCl_s \rightarrow K^+_{(aq)} + Cl^-_{(aq)}$	$CaCl_{2(s)} \rightarrow Ca^{2+}_{(aq)} + 2Cl^-_{(aq)}$
Amount of Cl^-	$0.266\text{ mol KCl} \times \frac{1\text{ mol } Cl^-}{1\text{ mol KCl}} = 0.266$ mol	$0.239\text{ mol } CaCl_2 \times \frac{2\text{ mol } Cl^-}{1\text{ mol } CaCl_2} = 0.478$ mol
Concentration of Cl^-	$\frac{0.266\text{ mol}}{0.100\text{ L}} = 2.66$ mol/L	$\frac{0.478\text{ mol}}{0.150\text{ L}} = 3.19$ mol/L

The concentration of chloride ions when 19.8 g of potassium chloride is dissolved in 100 mL of solution is 2.66 mol/L. The concentration of chloride ions when 26.5 g of calcium chloride is dissolved in 150 mL of solution is 3.19 mol/L.

(c) Total amount of $Cl^-_{(aq)} = 0.266 + 0.478$ mol
$= 0.744$ mol

Total volume of solution $= 0.100 + 0.150$ L
$= 0.250$ L

Total concentration of $Cl^-_{(aq)} = \frac{0.744\text{ mol}}{0.250\text{ L}}$
$= 2.98$ mol/L

The concentration of chloride ions when the solutions are mixed is 2.98 mol/L.

Check Your Solution

The units for amount and concentration are correct. The answers appear to be reasonable. When the solutions are mixed, the concentration of the chloride ions is not a simple average of the concentrations of the two solutions. Why? The volumes of the two solutions were different.

Sample Problem

The Mass Percent of Ions

Problem

The leaves of a rhubarb plant contain a relatively high concentration of oxalate ions, $C_2O_4^{2-}$. Oxalate ions are poisonous, causing respiratory failure. To determine the percent of oxalate ions, a student measured the mass of some leaves. Then the student ground up the leaves and added excess calcium chloride solution to precipitate calcium oxalate. The student tested 238.6 g of leaves. The dried mass of calcium oxalate was 0.556 g. What was the mass percent of oxalate ions in the leaves?

Although rhubarb stalks are safe to eat, the leaves are poisonous.

What Is Required?

You need to find the mass percent of oxalate ions in the leaves.

What Is Given?

You know the mass of the leaves is 238.6 g. You also know the mass of dried calcium oxalate is 0.556 g.

Plan Your Strategy

Determine the molar mass of calcium oxalate. Use the mass of calcium oxalate and its molar mass to find the amount (in mol) of calcium oxalate. Write the net ionic equation for the formation of calcium oxalate. From the coefficients in the net ionic equation, find the amount of oxalate ions (in mol). Calculate the mass of oxalate ions from the amount of oxalate ions (in mol) and the molar mass. Calculate the mass percent of oxalate ions in rhubarb leaves from the mass of the leaves and the mass of the oxalate ions present.

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Act on Your Strategy

Molar mass of CaC_2O_4 = 128.1 g

$$\text{Amount of } CaC_2O_4 = 0.556 \text{ g} \times \frac{1 \text{ mol}}{128.1 \text{ g}}$$
$$= 0.004\,34 \text{ mol}$$

The net ionic equation is

$$Ca^{2+}_{(aq)} + C_2O_4^{2-}{}_{(aq)} \rightarrow CaC_2O_{4(s)}$$

The mole ratio of CrO_4^{2-} to CaC_2O_4 is 1:1. Therefore, the leaves must have contained 0.004 34 mol of $C_2O_4^{2-}$.

Molar mass of $C_2O_4^{2-}$ = 88.02 g

$$\text{Mass of } C_2O_4^{2-} = 0.004\,34 \text{ mol} \times \frac{88.02 \text{ g}}{1 \text{ mol}}$$
$$= 0.382 \text{ g}$$

$$\text{Mass percent of } C_2O_4^{2-} \text{ in the leaves} = \frac{0.382 \text{ g}}{238.6 \text{ g}} \times 100\%$$
$$= 0.160\%$$

The mass percent of oxalate ions in the leaves is 0.160%.

Check Your Solution

Since the units divide out properly, you can be fairly confident that the answer is correct. The final value appears to be reasonable.

PROBLEM TIP

Review the method for solving stoichiometry problems you learned in Chapter 7, Section 7.1.

Sample Problem

Finding the Minimum Volume to Precipitate

Problem

Aqueous solutions that contain silver ions are usually treated with chloride ions to recover silver chloride. What is the minimum volume of 0.25 mol/L magnesium chloride, $MgCl_{2(aq)}$, needed to precipitate all the silver ions in 60 mL of 0.30 mol/L silver nitrate, $AgNO_{3(aq)}$? Assume that silver chloride is completely insoluble in water.

What Is Required?

You need to find the minimum volume of magnesium chloride that will precipitate all the silver ions.

What Is Given?

You know the volumes and concentrations of the silver nitrate (volume = 60 mL; concentration = 0.30 mol/L). The concentration of the magnesium chloride solution is 0.25 mol/L.

Plan Your Strategy

Find the amount (in mol) of silver nitrate from the volume and concentration of solution. Write a balanced chemical

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equation for the reaction. Use mole ratios from the coefficients in the equation to determine the amount (in mol) of magnesium chloride that is needed. Use the amount (in mol) of magnesium chloride and the concentration of solution to find the volume that is needed.

Act on Your Strategy

$$\text{Amount } AgNO_3 = 0.060 \cancel{L} \times 0.30 \text{ mol}/\cancel{L}$$
$$= 0.018 \text{ mol}$$

$$MgCl_{2(aq)} + 2AgNO_{3(aq)} \rightarrow 2AgCl_{(s)} + Mg(NO_3)_{2(aq)}$$

The mole ratio of $MgCl_2$ to $AgNO_3$ is 1:2.

$$n \text{ mol } MgCl_2 = 0.018 \text{ mol } \cancel{AgNO_3} \times \frac{1 \text{ mol } MgCl_2}{2 \text{ mol } \cancel{AgNO_3}}$$
$$= 0.0090 \text{ mol}$$

$$\text{Volume of 0.25 mol/L } MgCl_2 \text{ needed} = \frac{0.0090 \cancel{\text{ mol}}}{0.25 \cancel{\text{ mol}}/\text{L}}$$
$$= 0.036 \text{ L}$$

The minimum volume of 0.25 mol/L magnesium chloride that is needed is 36 mL.

Check Your Solution

The answer is in millilitres, an appropriate unit of volume. The amount appears to be reasonable.

Practice Problems

7. Food manufacturers sometimes add calcium acetate to puddings and sweet sauces as a thickening agent. What volume of 0.500 mol/L calcium acetate, $Ca(CH_3COO)_{2(aq)}$, contains 0.300 mol of acetate ions?

8. Ammonium phosphate can be used as a fertilizer. 6.0 g of ammonium phosphate is dissolved in sufficient water to produce 300 mL of solution. What are the concentrations (in mol/L) of the ammonium ions and phosphate ions present?

9. An aqueous solution of a certain salt contains chloride ions. A sample of this solution was made by dissolving 17.59 g of the salt in a 1 L volumetric flask. Then 25.00 mL of the solution was treated with excess silver nitrate. The precipitate, $AgCl_{(s)}$, was filtered and dried. If the mass of the dry precipitate was 47.35 g, what was the mass percent of chloride ions in the solution?

10. The active ingredient in some rat poisons is thallium(I) sulfate, Tl_2SO_4. A chemist takes a 500 mg sample of thallium(I) sulfate and adds potassium iodide, to precipitate yellow thallium(I) iodide. When the precipitate is dried, its mass is 200 mg. What is the mass percent of Tl_2SO_4 in the rat poison?

Limiting Reactant Problems in Aqueous Solutions

In Chapter 7, you learned how to solve limiting reactant problems. You can always recognize a limiting reactant problem because you are always given the amounts of both reactants. A key step in a limiting reactant problem is determining which one of the two reactants is limiting. In aqueous solutions, this usually means finding the amount of a reactant, given the volume and concentration of the solution.

Sample Problem

Finding the Mass of a Precipitated Compound

Problem

Mercury salts have a number of important uses in industry and in chemical analysis. Because mercury compounds are poisonous, however, the mercury ions must be removed from the waste water. Suppose that 25.00 mL of 0.085 mol/L aqueous sodium sulfide is added to 56.5 mL of 0.10 mol/L mercury(II) nitrate. What mass of mercury(II) sulfide, $HgS_{(s)}$, precipitates?

What Is Required?

You need to find the mass of mercury(II) sulfide that precipitates.

What Is Given?

You know the volumes and concentrations of the sodium sulfide and mercury(II) nitrate solutions.

Plan Your Strategy

Write a balanced chemical equation for the reaction. Find the amount (in mol) of each reactant, using its volume and concentration. Identify the limiting reactant. Determine the amount (in mol) of mercury(II) sulfide that forms. Calculate the mass of mercury(II) sulfide that precipitates.

Act on Your Strategy

The chemical equation is

$$Hg(NO_3)_{2(aq)} + Na_2S_{(aq)} \rightarrow 2NaNO_{3(aq)} + HgS_{(s)}$$

Calculate the amount (in mol) of each reactant.

$$\text{Amount of } Hg(NO_3)_2 = 0.0565 \cancel{L} \times 0.10 \text{ mol}/\cancel{L}$$
$$= 0.005\,65 \text{ mol}$$

$$\text{Amount of } Na_2S = 0.0250 \cancel{L} \times 0.085 \text{ mol}/\cancel{L}$$
$$= 0.002\,12 \text{ mol}$$

The reactants are in a 1:1 ratio. Because Na_2S is present in the smallest amount, it is the limiting reactant.
The equation indicates that each mol of Na_2S reacts to produce the same amount of $HgS_{(s)}$ precipitate. This amount is 0.002 12 mol.

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PROBLEM TIP

Chemists solve limiting reactant problems in different ways. The method used here is different from the one you used in Chapter 7, Section 7.2

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Molar mass of $HgS = 200.6 + 32.1$
$= 232.7$

Mass of $HgS_{(s)}$ = Amount × Molar mass

$$= 0.002\ 12\ \cancel{\text{mol}} \times \frac{232.7\ \text{g}}{1\ \cancel{\text{mol}}}$$

$= 0.493$ g

The mass of mercury(II) sulfide that precipitates is 0.49 g

Check Your Solution

The answer has appropriate units of mass. This answer appears to be reasonable, given the values in the problem.

Sample Problem

Finding the Mass of Another Precipitated Compound

Problem

Silver chromate, Ag_2CrO_4, is insoluble. It forms a brick-red precipitate. Calculate the mass of silver chromate that forms when 50.0 mL of 0.100 mol/L silver nitrate reacts with 25.0 mL of 0.150 mol/L sodium chromate.

What Is Required?

You need to find the mass of silver chromate that precipitates.

What Is Given?

You know the volumes and concentrations of the silver nitrate and sodium chromate solutions.

Plan Your Strategy

Write a balanced chemical equation for the reaction. Find the amount (in mol) of each reactant, using its volume and concentration. Identify the limiting reactant. Determine the amount (in mol) of silver chromate that forms. Calculate the mass of silver chromate that precipitates.

Act on Your Strategy

The chemical equation is

$$2AgNO_{3(aq)} + Na_2CrO_{4(aq)} \rightarrow Ag_2CrO_{4(s)} + 2NaNO_{3(aq)}$$

Calculate the amount (in mol) of each reactant.

Amount of $AgNO_3 = 0.0500\ \text{L} \times 0.100\ \text{mol/L}$
$= 5.00 \times 10^{-3}$ mol

Amount of $Na_2CrO_4 = 0.0250\ \text{L} \times 0.150\ \text{mol/L}$
$= 3.75 \times 10^{-3}$ mol

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To identify the limiting reactant, divide by the coefficient in the equation and find the smallest result.

$AgNO_3$: $5.00 \times \frac{10^{-3} \text{ mol}}{2} = 2.50 \times 10^{-3}$ mol

Na_2CrO_4: $3.75 \times \frac{10^{-3} \text{ mol}}{1} = 3.75 \times 10^{-3}$ mol

Since the smallest result is given for $AgNO_3$, this reactant is the limiting reactant.

Using the coefficients in the balanced equation, 2 mol of $AgNO_3$ react for each mole of Ag_2CrO_4 formed.

$$\text{Amount of } Ag_2CrO_4 = 5.00 \times 10^{-3} \text{ mol } AgNO_3 \times \frac{1 \text{ mol } Ag_2CrO_4}{2 \text{ mol } AgNO_3}$$
$$= 2.50 \times 10^{-3} \text{ mol } Ag_2CrO_4$$

The molar mass of Ag_2CrO_4 is 331.7 g/mol.

$$\text{Mass of precipitate} = 2.50 \times 10^{-3} \cancel{\text{mol}} \times \frac{331.7 \text{ g}}{1 \cancel{\text{mol}}}$$
$$= 0.829 \text{ g}$$

The mass of silver chromate that precipitates is 0.829 g.

Check Your Solution

The answer has appropriate units of mass. The answer appears to be reasonable, given the values in the problem.

Practice Problems

11. 8.76 g of sodium sulfide is added to 350 mL of 0.250 mol/L lead(II) nitrate solution. Calculate the maximum mass of precipitate that can form.

12. 25.0 mL of 0.400 mol/L $Pb(NO_3)_{2(aq)}$ is mixed with 300 mL of 0.220 mol/L $KI_{(aq)}$. What is the maximum mass of precipitate that can form?

13. A student mixes 15.0 mL of 0.250 mol/L aqueous sodium hydroxide with 20.0 mL of 0.400 mol/L aqueous aluminum nitrate.

(a) Write the chemical equation for the reaction.

(b) Calculate the maximum mass of precipitate that forms.

Section Wrap-up

In the last two sections, you have used qualitative and quantitative techniques to investigate ions in aqueous solution. Every drop of water that comes into your home contains a variety of such ions. It also contains other substances, in various concentrations. You have investigated water quality in previous grades. In the next section, you will consider the chemistry of water quality.

Section Review

1 I Equal volumes of 0.120 mol/L potassium nitrate and 0.160 mol/L iron(III) nitrate are mixed together. What is the concentration of nitrate ions in the mixture?

2 I Suppose that you want to remove the barium ions from 120 mL of 0.0500 mol/L aqueous barium nitrate solution. What is the minimum mass of sodium carbonate that you should add?

3 I An excess of aluminum foil is added to a certain volume of 0.675 mol/L aqueous copper(II) sulfate solution. The mass of solid copper that precipitates is measured and found to be 4.88 g. What was the volume of the copper(II) sulfate solution?

4 I To generate hydrogen gas, a student adds 5.77 g of mossy zinc to 80.1 mL of 4.00 mol/L hydrochloric acid in an Erlenmeyer flask. When the reaction is over, what is the concentration of aqueous zinc chloride in the flask?

5 I Copper can be recovered from scrap metal by adding sulfuric acid. Soluble copper sulfate is formed. The copper sulfate then reacts with metallic iron in a single displacement reaction. To simulate this reaction, a student places 1.942 g of iron wool in a beaker that contains 136.3 mL of 0.0750 mol/L aqueous copper(II) sulfate. What mass of copper is formed?

6 I Your stomach secretes hydrochloric acid to help you digest the food you have eaten. If too much HCl is secreted, however, you may need to take an antacid to neutralize the excess. One antacid product contains the compound magnesium hydroxide, $Mg(OH)_2$.

(a) Predict the reaction that takes place when magnesium hydroxide reacts with hydrochloric acid. (**Hint:** This is a double-displacement reaction.)

(b) Imagine that you are a chemical analyst testing the effectiveness of antacids. If 0.10 mol/L $HCl_{(aq)}$ serves as your model for stomach acid, how many litres will react with an antacid that contains 0.10 g of magnesium hydroxide?

7 I Even though lead is toxic, many lead compounds are still used as paint pigments (colourings). What volume of 1.50 mol/L lead(II) acetate contains 0.400 mol Pb^{2+} ions.