Differentiate between the following terms.

a. species reduced versus the reducing agent
b. species oxidized versus the oxidizing agent
c. oxidation state versus actual charge

22. How does one balance redox reactions by the oxidation states method?

Exercises

Aqueous Solutions: Strong and Weak Electrolytes

23. Show how each of the following strong electrolytes “breaks up” into its component ions upon dissolving in water by drawing molecular-level pictures.

a. NaBr
b. MgCl₂
c. Al₂(OH)₃

d. (NH₄)₂SO₄

e. NaOH

24. Match each name below with the following microscopic pictures of that compound in aqueous solution.

i. barium nitrate
ii. potassium carbonate
iii. sodium chloride
iv. magnesium sulfate

25. Calcium chloride is a strong electrolyte and is used as “salt” streets in the winter to melt ice and snow. Write a reaction to show how this substance breaks apart when it dissolves in water.

26. Commercial cold packs and hot packs are available for treating athletic injuries. Both types contain a pouch of water and a dry chemical. When the pack is struck, the pouch of water breaks, dissolving the chemical, and the solution becomes either hot or cold. Many hot packs use magnesium sulfate, and many cold packs use ammonium nitrate. Write reactions to show how these strong electrolytes break apart when they are added to a water.

27. A solution is prepared by dissolving 10.8 g of ammonium sulfate in 250.0 mL of water. What is the molarity of the solution?

28. A solution of ethanol (C₂H₅OH) in water is prepared by dissolving 75.0 mL of ethanol (density = 0.79 g/cm³) in enough water to make 250.0 mL of solution. What is the molarity of ethanol in this solution?
21. Differentiate between the following terms.
   a. species reduced versus the reducing agent
   b. species oxidized versus the oxidizing agent
   c. oxidation state versus actual charge
22. How does one balance redox reactions by the oxidation states method?

Exercises

In this section similar exercises are paired.

Aqueous Solutions: Strong and Weak Electrolytes

23. Show how each of the following strong electrolytes “breaks up” into its component ions upon dissolving in water by drawing molecular-level pictures.
   a. NaBr
   b. MgCl₂
   c. Al(NO₃)₃
   d. (NH₄)₂SO₄
   e. NaOH
24. Match each name below with the following microscopic pictures of that compound in aqueous solution.
   i. barium nitrate
   ii. potassium carbonate
   iii. sodium chloride
   iv. magnesium sulfate
   Which picture best represents HNO₃(aq)? Why aren’t any of the pictures a good representation of HC₂H₃O₂(aq)?
25. Calcium chloride is a strong electrolyte and is used to “salt” streets in the winter to melt ice and snow. Write a reaction to show how this substance breaks apart when it dissolves in water.
26. Commercial cold packs and hot packs are available for treating athletic injuries. Both types contain a pouch of water and a dry chemical. When the pack is struck, the pouch of water breaks, dissolving the chemical, and the solution becomes either hot or cold. Many hot packs use magnesium sulfate, and many cold packs use ammonium nitrate. Write reactions to show how these strong electrolytes break apart when they dissolve in water.

Solution Concentration: Molarity

27. Calculate the molarity of each of these solutions.
   a. A 5.623-g sample of NaHCO₃ is dissolved in enough water to make 250.0 mL of solution.
   b. A 184.6-mg sample of K₂Cr₂O₇ is dissolved in enough water to make 500.0 mL of solution.
   c. A 0.1025-g sample of copper metal is dissolved in 35 mL of concentrated HNO₃ to form Cu²⁺ ions and then water is added to make a total volume of 200.0 mL. (Calculate the molarity of Cu²⁺.)
28. A solution of ethanol (C₂H₅OH) in water is prepared by dissolving 75.0 mL of ethanol (density ~ 0.79 g/cm³) in enough water to make 250.0 mL of solution. What is the molarity of the ethanol in this solution?
29. Calculate the concentration of all ions present in each of the following solutions of strong electrolytes.
   a. 0.100 mole of Ca(NO₃)₂ in 100.0 mL of solution
   b. 2.50 moles of Na₂SO₄ in 1.25 L of solution
   c. 5.00 g of NH₄Cl in 500.0 mL of solution
   d. 1.00 g K₂PO₄ in 250.0 mL of solution
30. Calculate the concentration of all ions present in each of the following solutions of strong electrolytes.
   a. 0.0200 mole of sodium phosphate in 10.0 mL of solution
   b. 0.300 mole of barium nitrate in 600.0 mL of solution
   c. 1.00 g of potassium chloride in 0.500 L of solution
   d. 132 g of ammonium sulfate in 1.50 L of solution
31. Which of the following solutions of strong electrolytes contains the largest number of moles of chloride ions: 100.0 mL of 0.30 M AlCl₃, 50.0 mL of 0.60 M MgCl₂, or 200.0 mL of 0.40 M NaCl?
32. Which of the following solutions of strong electrolytes contains the largest number of ions: 100.0 mL of 0.100 M NaOH, 50.0 mL of 0.200 M BaCl₂, or 75.0 mL of 0.150 M Na₃PO₄?
33. What mass of NaOH is contained in 250.0 mL of a 0.400 M sodium hydroxide solution?
34. If 10.0 g of AgNO₃ is available, what volume of 0.25 M AgNO₃ solution can be prepared?
35. Describe how you would prepare 2.00 L of each of the following solutions.
   a. 0.250 M NaOH from solid NaOH
   b. 0.250 M NaOH from 1.00 M NaOH stock solution
   c. 0.100 M K₂Cr₂O₇ from solid K₂Cr₂O₇
   d. 0.100 M K₂Cr₂O₇ from 1.75 M K₂Cr₂O₇ stock solution
36. How would you prepare 1.00 L of a 0.50-M solution of each of the following?
   a. H₂SO₄ from “concentrated” (18 M) sulfuric acid
   b. HCl from “concentrated” (12 M) reagent
   c. NiCl₂ from the salt NiCl₂ · 6H₂O
   d. HNO₃ from “concentrated” (16 M) reagent
   e. Sodium carbonate from the pure solid
37. A solution is prepared by dissolving 10.8 g ammonium sulfate in enough water to make 100.0 mL of stock solution. A 10.00-mL sample of this stock solution is added to 50.00 mL of water. Calculate the concentration of ammonium ions and sulfate ions in the final solution.
38. A solution was prepared by mixing 50.00 mL of 0.100 M HNO₃ and 100.00 mL of 0.200 M HNO₃. Calculate the molarity of the final solution of nitric acid.
39. Calculate the sodium ion concentration when 70.0 mL of 3.0 M sodium carbonate is added to 30.0 mL of 1.0 M sodium bicarbonate.
40. Suppose 50.0 mL of 0.250 M CoCl₂ solution is added to 250.0 mL of 0.350 M NiCl₂ solution. Calculate the concentration, in moles per liter, of each of the ions present after mixing. Assume that the volumes are additive.

41. A standard solution is prepared for the analysis of fluoxymesterone (C₂₀H₂₉F₀₃), an anabolic steroid. A stock solution is first prepared by dissolving 1.584 g pure manganese metal in nitric acid and diluting to a final volume of 1.000 L. Calculate the concentration of the final solution in terms of molarity.

42. A stock solution containing Mn²⁺ ions was prepared by dissolving 1.584 g pure manganese metal in nitric acid and diluting to a final volume of 1.000 L. The following solutions were then prepared by dilution:

For solution A, 50.00 mL of stock solution was diluted to 1000.0 mL.

For solution B, 10.00 mL of solution A was diluted to 250.0 mL.

For solution C, 10.00 mL of solution B was diluted to 500.0 mL.

Calculate the concentrations of the stock solution and solutions A, B, and C.

Precipitation Reactions

43. On the basis of the general solubility rules given in Table 4.1, predict which of the following substances are likely to be soluble in water.

a. aluminum nitrate
b. magnesium chloride
c. rubidium sulfate
d. nickel(II) hydroxide
e. lead(II) sulfide
f. magnesium hydroxide
g. iron(III) phosphate

44. On the basis of the general solubility rules given in Table 4.1, predict which of the following substances are likely to be soluble in water.

a. zinc chloride
b. lead(II) nitrate
c. lead(II) sulfate
d. sodium iodide
e. cobalt(II) sulfide
f. chromium(III) hydroxide
g. magnesium carbonate
h. ammonium carbonate

45. When the following solutions are mixed together, what precipitate (if any) will form?

a. FeSO₄(aq) + KCl(aq)

b. Al(NO₃)₃(aq) + Ba(OH)₂(aq)
c. CaCl₂(aq) + Na₂SO₄(aq)
d. K₂S(aq) + Ni(NO₃)₂(aq)

46. When the following solutions are mixed together, what precipitate (if any) will form?

a. Hg₂(NO₃)₂(aq) + CuSO₄(aq)
b. Ni(NO₃)₂(aq) + CaCl₂(aq)
c. K₃CO₃(aq) + MgI₂(aq)
d. Na₂CrO₄(aq) + AlBr₃(aq)

47. For the reactions in Exercise 45, write the balanced formula equation, complete ionic equation, and net ionic equation. If no precipitate forms, write "No reaction."

48. For the reactions in Exercise 46, write the balanced formula equation, complete ionic equation, and net ionic equation. If no precipitate forms, write "No reaction."

49. Write the balanced formula and net ionic equation for the reaction that occurs when the contents of the two beakers are added together. What colors represent the spectator ions in each reaction?

a. 

b. 

c. 

50. Give an example how each of the following insoluble ionic compounds could be produced using a precipitation reaction. Write the balanced formula equation for each reaction.

a. Fe(OH)₃(s) c. PbSO₄(s)b. Hg₂Cl₂(s) d. BaCrO₄(s)

51. Write net ionic equations for the reaction, if any, that occur when aqueous solutions of the following are mixed.

a. ammonium sulfate and baryum nitrate
b. lead(II) nitrate and sodium chloride
c. sodium phosphate and potassium nitrate
d. sodium bromide and rubidium chloride
e. copper(II) chloride and sodium hydroxide

52. Write net ionic equations for the reaction, if any, that occur when aqueous solutions of the following are mixed.

a. chromium(III) chloride and sodium hydroxide
b. silver nitrate and ammonium carbonate
c. copper(II) sulfate and mercury(II) nitrate
d. strontium nitrate and potassium iodide
53. Separate samples of a solution of an unknown soluble ionic compound are treated with KCl, Na₂SO₄, and NaOH. A precipitate forms only when Na₂SO₄ is added. Which cations could be present in the unknown soluble ionic compound?

54. A sample may contain any or all of the following ions: Hg₂⁺, Ba²⁺, and Mn²⁺.
   a. No precipitate formed when an aqueous solution of NaCl was added to the sample solution.
   b. No precipitate formed when an aqueous solution of Na₂SO₄ was added to the sample solution.
   c. A precipitate formed when the sample solution was made basic with NaOH.

Which ion or ions are present in the sample solution?

55. What mass of Na₂CrO₄ is required to precipitate all of the silver ions from 75.0 mL of a 0.100-M solution of AgNO₃?

56. What volume of 0.100 M Na₂PO₄ is required to precipitate all the lead(II) ions from 150.0 mL of 0.250 M Pb(NO₃)₂?

57. What mass of solid aluminum hydroxide can be produced when 50.0 mL of 0.200 M Al(NO₃)₃ is added to 200.0 mL of 0.100 M KOH?

58. What mass of barium sulfate can be produced when 100.0 mL of a 0.100-M solution of barium chloride is mixed with 100.0 mL of a 0.100-M solution of iron(III) sulfate?

59. What mass of solid AgBr is produced when 100.0 mL of 0.130 M AgNO₃ is added to 20.0 mL of 1.00 M NaBr?

60. What mass of silver chloride can be prepared by the reaction of 100.0 mL of 0.20 M silver nitrate with 100.0 mL of 0.15 M calcium chloride? Calculate the concentrations of each ion remaining in solution after precipitation is complete.

61. A 100.0-mL aliquot of 0.200 M aqueous potassium hydroxide is mixed with 100.0 mL of 0.200 M aqueous magnesium nitrate.
   a. Write a balanced chemical equation for any reaction that occurs.
   b. What precipitate forms?
   c. What mass of precipitate is produced?
   d. Calculate the concentration of each ion remaining in solution after precipitation is complete.

62. The drawings below represent aqueous solutions. Solution A is 2.00 L of a 2.00-M aqueous solution of copper(II) nitrate. Solution B is 2.00 L of a 3.00-M aqueous solution of potassium hydroxide.

   a. Draw a picture of the solution made by mixing solutions A and B together after the precipitation reaction takes place. Make sure this picture shows the correct relative volume compared to solutions A and B, and the correct relative number of ions, along with the correct relative amount of solid formed.
   b. Determine the concentrations (in M) of all ions left in solution (from part a) and the mass of solid formed.

63. A 1.42-g sample of a pure compound, with formula M₂S₄O₆, was dissolved in water and treated with an excess of aqueous calcium chloride, resulting in the precipitation of all the sulfate ions as calcium sulfate. The precipitate was collected, dried, and found to weigh 1.36 g. Determine the atomic mass of M, and identify M.

64. You are given a 1.50-g mixture of sodium nitrate and sodium chloride. You dissolve this mixture into 100 mL of water and then add an excess of 0.500 M silver nitrate solution. You produce a white solid, which you then collect, dry, and measure. The white solid has a mass of 0.641 g.
   a. If you had an extremely magnified view of the solution (to the atomic-molecular level), list the species you would see (include charges, if any).
   b. Write the balanced net ionic equation for the reaction that produces the solid. Include phases and charges.
   c. Calculate the percent sodium chloride in the original unknown mixture.

**Acid–Base Reactions**

65. Write the balanced formula, complete ionic, and net ionic equations for each of the following acid–base reactions.
   a. HClO₄(aq) + Mg(OH)₂(s) →
   b. HCN(aq) + NaOH(aq) →
   c. HCl(aq) + NaOH(aq) →

66. Write the balanced formula, complete ionic, and net ionic equations for each of the following acid–base reactions.
   a. HNO₃(aq) + Al(OH)₃(s) →
   b. HC₂H₃O₂(aq) + KOH(aq) →
   c. Ca(OH)₂(aq) + HCl(aq) →

67. Write the balanced formula equation for the acid–base reactions that occur when the following are mixed.
   a. potassium hydroxide (aqueous) and nitric acid
   b. barium hydroxide (aqueous) and hydrochloric acid
   c. perchloric acid [HClO₄(aq)] and solid iron(III) hydroxide
   d. solid silver hydroxide and hydrobromic acid
   e. aqueous strontium hydroxide and hydroiodic acid

68. What acid and what base would react in aqueous solution so that the following salts appear as products in the formula equation? Write the balanced formula equation for each reaction.
   a. potassium perchlorate
   b. cesium nitrate
   c. calcium iodate

69. What volume of each of the following solutions will react completely with 50.00 mL of 0.200 M NaOH?
   a. 0.100 M HCl
   b. 0.150 M HNO₃
   c. 0.200 M HC₂H₃O₂ (1 acidic hydrogen)

70. What volume of each of the following solutions will react completely with 25.00 mL of 0.200 M HCl?
   a. 0.100 M NaOH
   b. 0.0500 M Sr(OH)₂
   c. 0.250 M KOH
71. Hydrochloric acid (75.0 mL of 0.250 M) is added to 225.0 mL of 0.0550 M Ba(OH)\(_2\) solution. What is the concentration of the excess H\(^+\) or OH\(^-\) ions left in this solution?

72. A student mixes four reagents together, thinking that the solutions will neutralize each other. The solutions mixed together are 50.0 mL of 0.100 M hydrochloric acid, 100.0 mL of 0.200 M nitric acid, 500.0 mL of 0.010 M calcium hydroxide, and 200.0 mL of 0.100 M rubidium hydroxide. Did the acids and bases exactly neutralize each other? If not, calculate the concentration of excess H\(^+\) or OH\(^-\) ions left in solution.

73. A 25.00-mL sample of hydrochloric acid solution requires 24.16 mL of 0.106 M sodium hydroxide solution to neutralize it. What is the molarity of the hydrochloric acid in this solution?

74. A 10.00-mL sample of vinegar, an aqueous solution of acetic acid (HC\(_2\)H\(_3\)O\(_2\)), is titrated with 0.5062 M NaOH and 16.58 mL is required to reach the equivalence point.
   a. What is the molarity of the acetic acid?
   b. If the density of the vinegar is 1.006 g/cm\(^3\), what is the percent mass of acetic acid in the vinegar?
   c. What volume of 0.0200 M calcium hydroxide is required to neutralize 35.00 mL of 0.0500 M nitric acid?

75. A 30.0-mL sample of an unknown strong base is neutralized after the addition of 12.0 mL of a 0.150 M HNO\(_3\) solution. If the unknown base concentration is 0.0300 M, give some possible identities for the unknown base.

76. A student titrates an unknown amount of potassium hydroxide phthalate (KHC\(_8\)H\(_4\)O\(_4\)), often abbreviated KHP, with 0.150 M HNO\(_3\) solution. KHP (molar mass = 204.22 g/mol) has one acidic hydrogen. What mass of KHP was titrated (reacted completely) by the sodium hydroxide solution?

77. The concentration of a certain sodium hydroxide solution was determined by using the solution to titrate a sample of potassium hydrogen phthalate (abbreviated as KHP). KHP is an acid with one acidic hydrogen and a molar mass of 204.22 g/mol. In the titration, 34.67 mL of the sodium hydroxide solution was required to react with 0.1082 g KHP. Calculate the molarity of the sodium hydroxide.

### Oxidation–Reduction Reactions

79. Assign oxidation states for all atoms in each of the following compounds.
   a. KMnO\(_4\)
   b. NiO\(_2\)
   c. Na\(_2\)Fe(OH)\(_6\)
   d. (NH\(_4\))\(_2\)HPO\(_4\)
   e. PO\(_4\)
   f. Fe\(_2\)O\(_4\)
   g. XeOF\(_4\)
   h. SF\(_4\)
   i. CO
   j. C\(_6\)H\(_5\)O\(_6\)

80. Assign oxidation states for all atoms in each of the following compounds.
   a. UO\(_2^{2+}\)
   b. As\(_2\)O\(_3\)
   c. NaBiO\(_3\)
   d. Na\(_2\)AsO\(_3\)
   e. H\(_2\)AsO\(_4\)
   f. Mg\(_2\)P\(_2\)O\(_7\)
   g. Na\(_2\)S\(_2\)O\(_3\)
   h. HgCl\(_2\)
   i. Ca(NO\(_3\))\(_2\)

81. Assign the oxidation state for nitrogen in each of the following.
   a. Li\(_3\)N
   b. NH\(_3\)
   c. N\(_2\)H\(_4\)
   d. NO
   e. N\(_2\)O
   f. NO\(_3\)
   g. NO\(_2\)
   h. NO\(_2\)⁻
   i. N\(_2\)
   j. N\(_2\)O
   k. N\(_2\)O\(_3\)

82. Assign oxidation numbers to all the atoms in each of the following.
   a. SrCr\(_2\)O\(_7\)
   b. CuCl\(_2\)
   c. O\(_2\)
   d. H\(_2\)O\(_2\)
   e. MgCO\(_3\)
   f. (NH\(_2\))\(_2\)Ce(SO\(_4\))\(_3\)
   g. PbO\(_2\)

83. Specify which of the following are oxidation–reduction reactions, and identify the oxidizing agent, the reducing agent, the substance being oxidized, and the substance being reduced.
   a. Cu(s) + 2Ag\(^+\)(aq) → 2Ag(s) + Cu\(^{2+}\)(aq)
   b. HCl(g) + NH\(_3\)(g) → NH\(_4\)Cl(s)
   c. SiCl\(_4\)(l) + H\(_2\)O(l) → 4HCl(aq) + SiO\(_2\)(s)
   d. SiCl\(_4\)(l) + 2Mg(s) → 2MgCl\(_2\)(s) + Si(s)
   e. Al(OH)\(_3\)(aq) → AlO\(_2\)(aq) + 2H\(_2\)O(l)

84. Specify which of the following equations represent oxidation–reduction reactions, and indicate the oxidizing agent, the reducing agent, the species being oxidized, and the species being reduced.
   a. CH\(_4\)(g) + H\(_2\)O(g) → CO(g) + 3H\(_2\)(g)
   b. 2AgNO\(_3\)(aq) + Cu(s) → Cu(NO\(_3\))\(_2\)(aq) + 2Ag(s)
   c. Zn(s) + 2HCl(aq) → ZnCl\(_2\)(aq) + H\(_2\)(g)
   d. 2H\(^+\)(aq) + 2CrO\(_4\)\(^{2-}\)(aq) → Cr\(_2\)O\(_7\)\(^{2-}\)(aq) + H\(_2\)O(l)

85. Consider the reaction between sodium metal and fluorine (F\(_2\)) gas to form sodium fluoride. Using oxidation states, how many electrons would each sodium atom lose, and how many electrons would each fluorine atom gain? How many sodium atoms are needed to react with one fluorine molecule? Write a balanced equation for this reaction.

86. Consider the reaction between oxygen (O\(_2\)) gas and magnesium metal to form magnesium oxide. Using oxidation states, how many electrons would each magnesium atom lose, and how many electrons would each oxygen atom gain? How many magnesium atoms are needed to react with one oxygen molecule? Write a balanced equation for this reaction.

87. Balance each of the following oxidation–reduction reactions by using the oxidation states method.
   a. C\(_2\)H\(_2\)(g) + O\(_2\)(g) → CO\(_2\)(g) + H\(_2\)O(g)
   b. Mg(s) + HCl(aq) → Mg\(^{2+}\)(aq) + Cl\(^-\)(aq) + H\(_2\)(g)
   c. Co\(^{3+}\)(aq) + Ni(s) → Co\(^{2+}\)(aq) + Ni\(^{2+}\)(aq)
   d. Zn(s) + H\(_2\)SO\(_4\)(aq) → ZnSO\(_4\)(aq) + H\(_2\)(g)

88. Balance each of the following oxidation–reduction reactions by using the oxidation states method.
   a. Cl\(_2\)(g) + Al(s) → Al\(^{3+}\)(aq) + Cl\(^-\)(aq)
   b. O\(_2\)(g) + H\(_2\)O(l) + Pb(s) → Pb(OH)\(_2\)(s)
   c. H\(^+\)(aq) + MnO\(_4\)\(^{-}\)(aq) + Fe\(^{2+}\)(aq) → Mn\(^{2+}\)(aq) + Fe\(^{3+}\)(aq) + H\(_2\)O(l)
2. What are the oxidation states for each atom?

\[
\begin{align*}
\text{PbO} & \quad +2 \\
\text{NH}_3 & \quad +1 \\
\rightarrow & \\
\text{N}_2 & \quad 0 \\
\text{H}_2\text{O} & \quad +1 \\
\text{Pb} & \quad 0
\end{align*}
\]

3. How are electrons gained and lost?

\[
\begin{align*}
\text{PbO} & \quad +2 \\
\text{NH}_3 & \quad +1 \\
\rightarrow & \\
\text{N}_2 & \quad 0 \\
\text{H}_2\text{O} & \quad +1 \\
\text{Pb} & \quad 0
\end{align*}
\]

The oxidation states of all other atoms are unchanged.

4. What coefficients are needed to equalize the electrons gained and lost?

\[
\begin{align*}
\text{PbO} & \quad +2 \\
\text{NH}_3 & \quad +1 \\
\rightarrow & \\
\text{N}_2 & \quad 0 \\
\text{H}_2\text{O} & \quad +1 \\
\text{Pb} & \quad 0
\end{align*}
\]

5. What coefficients are needed to balance the remaining elements?

Balance O:

\[
\begin{align*}
3\text{PbO} + 2\text{NH}_3 & \rightarrow \text{N}_2 + 3\text{H}_2\text{O} + 3\text{Pb}
\end{align*}
\]

All the elements are now balanced. The balanced equation with states is:

\[
3\text{PbO}(s) + 2\text{NH}_3(g) \rightarrow \text{N}_2(g) + 3\text{H}_2\text{O}(l) + 3\text{Pb}(s)
\]

See Exercises 4.87 and 4.88

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**Key terms**

- aqueous solution
- polar molecule
- hydration
- solubility
- solute
- solvent
- electrical conductivity
- strong electrolyte
- weak electrolyte
- nonelectrolyte
- acid
- strong acid
- strong base

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**Chemical reactions in solution are very important in everyday life.**

**Water is a polar solvent that dissolves many ionic and polar substances.**

**Electrolytes**

- Strong electrolyte: 100% dissociated to produce separate ions; strongly conducts an electric current
- Weak electrolyte: Only a small percentage of dissolved molecules produce ions; weakly conducts an electric current
- Nonelectrolyte: Dissolved substance produces no ions; does not conduct an electric current

**Acids and bases**

- Arrhenius model
  - Acid: produces H⁺
  - Base: produces OH⁻
Chapter 4  Types of Chemical Reactions and Solution Stoichiometry

Key terms
- weak acid
- weak base

Section 4.3
- molarity
- standard solution
- dilution

Section 4.5
- precipitation reaction
- precipitate

Section 4.6
- formula equation
- complete ionic equation
- spectator ions
- net ionic equation

Section 4.8
- acid
- base
- neutralization reaction
- volumetric analysis
- titration
- stoichiometric (equivalence) point
- indicator
- endpoint

Section 4.9
- oxidation–reduction (redox) reaction
- oxidation state
- oxidation
- reduction
- oxidizing agent (electron acceptor)
- reducing agent (electron donor)

Acids and bases
- Bronsted-Lowry model
  - Acid: proton donor
  - Base: proton acceptor
- Strong acid: completely dissociates into separated $H^+$ and anions
- Weak acid: dissociates to a slight extent

Molarity
- One way to describe solution composition
  \[ \text{Molarity} (M) = \frac{\text{moles of solute}}{\text{volume of solution} (L)} \]
- Moles solute = volume of solution (L) $\times$ molarity
- Standard solution: molarity is accurately known

Dilution
- Solvent is added to reduce the molarity
- Moles of solute after dilution = moles of solute before dilution \[ M_1V_1 = M_2V_2 \]

Types of equations that describe solution reactions
- Formula equation: All reactants and products are written as complete formulas
- Complete ionic equation: All reactants and products that are strong electrolytes are written as separated ions
- Net ionic equation: Only those compounds that undergo a change are written; spectator ions are not included

Solubility rules
- Based on experiment observation
- Help predict the outcomes of precipitation reactions

Important types of solution reactions
- Acid–base reactions: involve a transfer of $H^+$ ions
- Precipitation reactions: formation of a solid occurs
- Oxidation–reduction reactions: involve electron transfer

Titrations
- Measures the volume of a standard solution (titrant) needed to react with a substance in solution
- Stoichiometric (equivalence) point: the point at which the required amount of titrant has been added to exactly react with the substance being analyzed
- Endpoint: the point at which a chemical indicator changes color
Oxidation–reduction reactions

- Oxidation states are assigned using a set of rules to keep track of electron flow.
- Oxidation: increase in oxidation state (a loss of electrons)
- Reduction: decrease in oxidation state (a gain of electrons)
- Oxidizing agent: gains electrons (is reduced)
- Reducing agent: loses electrons (is oxidized)
- Equations for oxidation–reduction reactions can be balanced by the oxidation states method.

Review questions

1. The \((aq)\) designation listed after a solute indicates the process of hydration. Using \(\text{KBr}(aq)\) and \(\text{C}_2\text{H}_5\text{OH}(aq)\) as your examples, explain the process of hydration for soluble ionic compounds and for soluble covalent compounds.

2. Characterize strong electrolytes versus weak electrolytes versus nonelectrolytes. Give examples of each. How do you experimentally determine whether a soluble substance is a strong electrolyte, weak electrolyte, or nonelectrolyte?

3. Distinguish between the terms slightly soluble and weak electrolyte.

4. Molarity is a conversion factor relating moles of solute in solution to the volume of the solution. How does one use molarity as a conversion factor to convert from moles of solute to volume of solution, and from volume of solution to moles of solute present?

5. What is a dilution? What stays constant in a dilution? Explain why the equation \(M_1 V_1 = M_2 V_2\) works for dilution problems.

6. When the following beakers are mixed, draw a molecular-level representation of the product mixture (see Fig. 4.17).

7. Differentiate between the formula equation, the complete ionic equation, and the net ionic equation. For each reaction in Question 6, write all three balanced equations.

8. What is an acid–base reaction? Strong bases are soluble ionic compounds that contain the hydroxide ion. List the strong bases. When a strong base reacts with an acid, what is always produced? Explain the terms titration, stoichiometric point, neutralization, and standardization.

9. Define the terms oxidation, reduction, oxidizing agent, and reducing agent. Given a chemical reaction, how can you tell if it is a redox reaction?

10. Consider the steps involved in balancing oxidation–reduction reactions by using oxidation states. The key to the oxidation states method is to equalize the electrons lost by the species oxidized with the electrons gained by the species reduced. First of all, how do you recognize what is oxidized and what is reduced? Second, how do you balance the electrons lost with the electrons gained? Once the electrons are balanced, what else is needed to balance the oxidation–reduction reaction?

Active Learning Questions

These questions are designed to be used by groups of students in class.

1. Assume you have a highly magnified view of a solution of \(\text{HCl}\) that allows you to "see" the \(\text{HCl}\). Draw this magnified view. If you dropped in a piece of magnesium, the magnesium would disappear and hydrogen gas would be released. Represent this change using symbols for the elements, and write out the balanced equation.

2. You have a solution of table salt in water. What happens to the salt concentration (increases, decreases, or stays the same) as the solution boils? Draw pictures to explain your answer.
Note that hydrogen, chlorine, and oxygen do not change oxidation states and are not involved in electron exchange. Thus we focus our attention on Sn and N:

\[
\begin{align*}
\text{Sn} & \quad + \quad \text{NO}_3^- & \rightarrow & \quad \text{SnCl}_6^{2-} & + & \quad \text{NO}_2 \\
+5 & & & +4 & & +4
\end{align*}
\]

This means we need a coefficient of 4 for the N-containing species.

\[H^+ + Cl^- + \text{Sn} + 4\text{NO}_3^- \rightarrow \text{SnCl}_6^{2-} + 4\text{NO}_2 + \text{H}_2\text{O}\]

Now we balance the rest of the equation by inspection.

Balance Cl\(-\):

\[H^+ + 6\text{Cl}^- + \text{Sn} + 4\text{NO}_3^- \rightarrow \text{SnCl}_6^{2-} + 4\text{NO}_2 + \text{H}_2\text{O}\]

Balance O:

\[H^+ + 6\text{Cl}^- + \text{Sn} + 4\text{NO}_3^- \rightarrow \text{SnCl}_6^{2-} + 4\text{NO}_2 + 4\text{H}_2\text{O}\]

Balance H:

\[8H^+ + 6\text{Cl}^- + \text{Sn} + 4\text{NO}_3^- \rightarrow \text{SnCl}_6^{2-} + 4\text{NO}_2 + 4\text{H}_2\text{O}\]

This gives the final balanced equation. Now we will write the equation with the states included:

\[8H^+(aq) + 6\text{Cl}^-(aq) + \text{Sn}(s) + 4\text{NO}_3^-(aq) \rightarrow \text{SnCl}_6^{2-}(aq) + 4\text{NO}_2(g) + 4\text{H}_2\text{O}(l)\]

### Problem-Solving Strategy

**Balancing Oxidation–Reduction Reactions by Oxidation States**

1. Write the unbalanced equation.
2. Determine the oxidation states of all atoms in the reactants and products.
3. Show electrons gained and lost using “tie lines.”
4. Use coefficients to equalize the electrons gained and lost.
5. Balance the rest of the equation by inspection.
6. Add appropriate states.

### Example 4.18

**Balancing Oxidation–Reduction Reactions**

Balance the reaction between solid lead(II) oxide and ammonia gas to produce nitrogen gas, liquid water, and solid lead.

**Solution**

We'll use the Problem-Solving Strategy for Balancing Oxidation–Reduction Reactions by Oxidation States.

1. **What is the unbalanced equation?**

\[\text{PbO}(s) + \text{NH}_3(g) \rightarrow \text{N}_2(g) + \text{H}_2\text{O}(l) + \text{Pb}(s)\]
Lead is reduced (its oxidation state decreases from +2 to 0), and carbon is oxidized (its oxidation state increases from +2 to +4). PbO is the oxidizing agent, and CO is the reducing agent.

Critical Thinking
Dalton believed that atoms were indivisible. Thomson and Rutherford helped to show that this was not true. What if atoms were indivisible? How would this affect the types of reactions you have learned about in this chapter?

4.10 Balancing Oxidation–Reduction Equations

It is important to be able to balance oxidation–reduction reactions. One method involves the use of oxidation states (discussed in this section), and the other method (normally used for more complex reactions) involves separating the reaction into two half-reactions. We’ll discuss the second method for balancing oxidation–reduction reactions in Chapter 18.

Oxidation States Method of Balancing Oxidation–Reduction Reactions

Consider the reaction between solid copper and silver ions in aqueous solution:

\[ \text{Cu(s)} + \text{Ag}^+(aq) \rightarrow \text{Ag(s)} + \text{Cu}^{2+}(aq) \]

We can tell this is a redox reaction by assigning oxidation states as follows:

\[
\begin{align*}
\text{Cu} &: 0 \\
\text{Ag}^+ &: +1 \\
\text{Ag} &: 0 \\
\text{Cu}^{2+} &: +2 \\
\end{align*}
\]

We know that in an oxidation–reduction reaction we must ultimately have equal numbers of electrons gained and lost, and we can use this principle to balance redox equations. For example, in this case, 2 Ag$^+$ ions must be reduced for every Cu atom oxidized:

\[ \text{Cu(s)} + 2\text{Ag}^+(aq) \rightarrow 2\text{Ag(s)} + \text{Cu}^{2+}(aq) \]

This gives us the balanced equation.

Now consider a more complex reaction:

\[ \text{H}^+(aq) + \text{Cl}^-(aq) + \text{Sn(s)} + \text{NO}_3^-(aq) \rightarrow \text{SnCl}_2^-(aq) + \text{NO}_2(g) + \text{H}_2\text{O(l)} \]

To balance this equation by oxidation states, we first need to assign the oxidation states to all the atoms in the reactants and products.

\[
\begin{align*}
\text{H}^+ &: +1 \\
\text{Cl}^- &: -1 \\
\text{Sn} &: +5 \\
\text{NO}_3^- &: -2 \\
\text{SnCl}_2^- &: +4 \\
\text{NO}_2 &: +4 \\
\text{H}_2\text{O} &: -2 \\
\end{align*}
\]
With this background, we can now define some important terms. **Oxidation** is an increase in oxidation state (a loss of electrons). **Reduction** is a decrease in oxidation state (a gain of electrons). Thus in the reaction

\[
2\text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s)
\]

sodium is oxidized and chlorine is reduced. In addition, Cl\(_2\) is called the *oxidizing agent* (electron acceptor), and Na is called the *reducing agent* (electron donor). These terms are summarized in Fig. 4.20.

Concerning the reaction

\[
\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g)
\]

we can say the following:

- Methane is oxidized because there has been an increase in carbon’s oxidation state (the carbon atom has formally lost electrons).
- Oxygen is reduced because there has been a decrease in its oxidation state (oxygen has formally gained electrons).
- CH\(_4\) is the reducing agent.
- O\(_2\) is the oxidizing agent.

Note that when the oxidizing or reducing agent is named, the *whole compound* is specified, not just the element that undergoes the change in oxidation state.

### Oxidation–Reduction Reactions

Metallurgy, the process of producing a metal from its ore, always involves oxidation–reduction reactions. In the metallurgy of galena (PbS), the principal lead-containing ore, the first step is the conversion of lead sulfide to its oxide (a process called roasting):

\[
2\text{PbS}(s) + 3\text{O}_2(g) \rightarrow 2\text{PbO}(s) + 2\text{SO}_2(g)
\]

The oxide is then treated with carbon monoxide to produce the free metal:

\[
\text{PbO}(s) + \text{CO}(g) \rightarrow \text{Pb}(s) + \text{CO}_2(g)
\]

For each reaction, identify the atoms that are oxidized and reduced, and specify the oxidizing and reducing agents.

**Solution**

For the first reaction, we can assign the following oxidation states:

\[
2\text{PbS}(s) + 3\text{O}_2(g) \rightarrow 2\text{PbO}(s) + 2\text{SO}_2(g)
\]

The oxidation state for the sulfur atom increases from -2 to +4. Thus sulfur is oxidized. The oxidation state for each oxygen atom decreases from 0 to -2. Oxygen is reduced. The oxidizing agent (that accepts the electrons) is O\(_2\), and the reducing agent (that donates electrons) is PbS.

For the second reaction we have

\[
\text{PbO}(s) + \text{CO}(g) \rightarrow \text{Pb}(s) + \text{CO}_2(g)
\]

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**Interactive Example 4.17**

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with the sodium hydroxide solution to the phenolphthalein endpoint. The difference between the final and initial buret readings indicates that 41.20 mL of the sodium hydroxide solution is required to react exactly with the 1.3009 g KHP. Calculate the concentration of the sodium hydroxide solution.

Solution

Where are we going?
To find the concentration of NaOH solution

What do we know?
1. 1.3009 g KHC$_8$H$_4$O$_4$ (KHP), molar mass (204.22 g/mol)
2. 41.20 mL NaOH solution to neutralize KHP
3. The chemical reaction
   \[
   \text{HC}_8\text{H}_4\text{O}_4^-(aq) + \text{OH}^- (aq) \rightarrow \text{H}_2\text{O}(l) + \text{C}_8\text{H}_4\text{O}_4^{2-}(aq)
   \]

How do we get there?
Use the Problem-Solving Strategy for Performing Calculations for Acid–Base Reactions.
1. What are the ions present in the combined solution?
   \[
   \text{K}^+ \quad \text{HC}_8\text{H}_4\text{O}_4^- \quad \text{Na}^+ \quad \text{OH}^-
   \]
2. What is the balanced net ionic equation for the reaction?
   \[
   \text{HC}_8\text{H}_4\text{O}_4^- (aq) + \text{OH}^- (aq) \rightarrow \text{H}_2\text{O}(l) + \text{C}_8\text{H}_4\text{O}_4^{2-}(aq)
   \]
3. What are the moles of KHP?
   \[
   1.3009 \text{ g KHC}_8\text{H}_4\text{O}_4 \times \frac{1 \text{ mol KHC}_8\text{H}_4\text{O}_4}{204.22 \text{ g KHC}_8\text{H}_4\text{O}_4} = 6.3701 \times 10^{-3} \text{ mol KHC}_8\text{H}_4\text{O}_4
   \]
4. Which reactant is limiting?
   This problem requires the addition of just enough OH$^-$ ions to react exactly with the KHP present. We do not need to be concerned with limiting reactant here.
5. What moles of OH$^-$ are required?
   6.3701 \times 10^{-3} \text{ mole of OH}^- is required to neutralize the KHP present.
6. What is the molarity of the NaOH solution?
   \[
   \text{Molarity of NaOH} = \frac{\text{mol NaOH}}{\text{L solution}} = \frac{6.3701 \times 10^{-3} \text{ mol NaOH}}{4.120 \times 10^{-2} \text{ L}} = 0.1546 \text{ M}
   \]
This standard sodium hydroxide solution can now be used in other experiments (see Example 4.15).

Critical Thinking

In Example 4.14 you determined the concentration of an aqueous solution of NaOH using phenolphthalein as an indicator. What if you used an indicator for which the endpoint of the titration occurs after the equivalence point? How would this affect your calculated concentration of NaOH?
When the analyte is a base or an acid, the required titrant is a strong acid or strong base, respectively. This procedure is called an acid–base titration. An indicator very commonly used for acid–base titrations is phenolphthalein, which is colorless in an acidic solution and pink in a basic solution. Thus, when an acid is titrated with a base, the phenolphthalein remains colorless until after the acid is consumed and the first drop of excess base is added. In this case, the endpoint (the solution changes from colorless to pink) occurs approximately one drop of base beyond the stoichiometric point. This type of titration is illustrated in Fig. 4.18.

We will deal with the acid–base titrations only briefly here but will return to the topic of titrations and indicators in more detail in Chapter 15. The titration of an acid with a standard solution containing hydroxide ions is described in Example 4.15. In Example 4.14 we show how to determine accurately the concentration of a sodium hydroxide solution. This procedure is called standardizing the solution.

Neutralization Titration

A student carries out an experiment to standardize (determine the exact concentration of) a sodium hydroxide solution. To do this, the student weighs out a 1.3009-g sample of potassium hydrogen phthalate (KHC₈H₄O₄, often abbreviated KHP). KHP (molar mass 204.22 g/mol) has one acidic hydrogen. The student dissolves the KHP in distilled water, adds phenolphthalein as an indicator, and titrates the resulting solution

Interactive Example 4.14

Sign in at [http://login.cengagebrain.com](http://login.cengagebrain.com) to try this Interactive Example in OWL.
Determing the Mass of Product Formed II

When aqueous solutions of Na₂SO₄ and Pb(NO₃)₂ are mixed, PbSO₄ precipitates. Calculate the mass of PbSO₄ formed when 1.25 L of 0.0500 M Pb(NO₃)₂ and 2.00 L of 0.0250 M Na₂SO₄ are mixed.

Solution

Where are we going?
To find the mass of solid PbSO₄ formed

What do we know?
  › 1.25 L of 0.0500 M Pb(NO₃)₂
  › 2.00 L of 0.0250 M Na₂SO₄
  › Chemical reaction

\[
Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)
\]

What information do we need?
  › The limiting reactant

How do we get there?
1. What are the ions present in the combined solution?

\[
Na^+ \quad SO_4^{2-} \quad Pb^{2+} \quad NO_3^{-}
\]

What is the reaction?
Since NaN0₃ is soluble and PbSO₄ is insoluble, solid PbSO₄ will form.

2. What is the balanced net ionic equation for the reaction?

\[
Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)
\]

3. What are the moles of reactants present in the solution?

\[
1.25 \text{ L} \times \frac{0.0500 \text{ mol } Pb^{2+}}{\text{L}} = 0.0625 \text{ mol } Pb^{2+}
\]

\[
2.00 \text{ L} \times \frac{0.0250 \text{ mol } SO_4^{2-}}{\text{L}} = 0.0500 \text{ mol } SO_4^{2-}
\]

4. Which reactant is limiting?
Because Pb²⁺ and SO₄²⁻ react in a 1:1 ratio, the amount of SO₄²⁻ will be limiting (0.0500 mol SO₄²⁻ is less than 0.0625 mole of Pb²⁺).

5. What number of moles of PbSO₄ will be formed?
Since SO₄²⁻ is limiting, only 0.0500 mole of solid PbSO₄ will be formed.

6. What mass of PbSO₄ will be formed?

\[
0.0500 \text{ mol PbSO}_4 \times \frac{303.3 \text{ g PbSO}_4}{1 \text{ mol PbSO}_4} = 15.2 \text{ g PbSO}_4
\]

See Exercises 4.57 and 4.58
4.7 Stoichiometry of Precipitation Reactions

Solution

Where are we going?
To find the mass of solid NaCl required to precipitate the Ag+

What do we know?
- 1.50 L of 0.100 M AgNO₃

What information do we need to find the mass of NaCl?
- Moles of Ag⁺ in the solution

How do we get there?
What are the ions present in the combined solution?

\[
\text{Ag}^+ \quad \text{NO}_3^- \quad \text{Na}^+ \quad \text{Cl}^-
\]

What is the balanced net ionic equation for the reaction?
Note from Table 4.1 that NaN₀₃ is soluble and that AgCl is insoluble. Therefore, solid AgCl forms according to the following net ionic equation:

\[
\text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl(s)}
\]

What are the moles of Ag⁺ ions present in the solution?

\[
1.50 \text{ L} \times \frac{0.100 \text{ mol Ag}^+}{\text{L}} = 0.150 \text{ mol Ag}^+
\]

How many moles of Cl⁻ are required to react with all the Ag⁺?
Because Ag⁺ and Cl⁻ react in a 1:1 ratio, 0.150 mole of Cl⁻ and thus 0.150 mole of NaCl are required.

What mass of NaCl is required?

\[
0.150 \text{ mol NaCl} \times \frac{58.44 \text{ g NaCl}}{\text{mol NaCl}} = 8.77 \text{ g NaCl}
\]

See Exercise 4.55

Notice from Example 4.19 that the procedures for doing stoichiometric calculations for solution reactions are very similar to those for other types of reactions. It is useful to think in terms of the following steps for reactions in solution.

**Problem-Solving Strategy**

Solving Stoichiometry Problems for Reactions in Solution

1. Identify the species present in the combined solution, and determine what reaction occurs.
2. Write the balanced net ionic equation for the reaction.
3. Calculate the moles of reactants.
4. Determine which reactant is limiting.
5. Calculate the moles of product or products, as required.
6. Convert to grams or other units, as required.
Canceling the spectator ions

\[ \text{K}^+(aq) + \text{Cl}^-(aq) + \text{Ag}^+(aq) + \text{NO}_3^-(aq) \rightarrow \text{AgCl}(s) + \text{K}^+(aq) + \text{NO}_3^-(aq) \]

gives the following net ionic equation.

**Net Ionic Equation**

\[ \text{Cl}^-(aq) + \text{Ag}^+(aq) \rightarrow \text{AgCl}(s) \]

b. **Formula Equation**

\[ 3\text{KOH}(aq) + \text{Fe(NO}_3)_3(aq) \rightarrow \text{Fe(OH)}_3(s) + 3\text{KNO}_3(aq) \]

**Complete Ionic Equation**

\[ 3\text{K}^+(aq) + 3\text{OH}^-(aq) + \text{Fe}^{3+}(aq) + 3\text{NO}_3^-(aq) \rightarrow \text{Fe(OH)}_3(s) + 3\text{K}^+(aq) + 3\text{NO}_3^-(aq) \]

**Net Ionic Equation**

\[ 3\text{OH}^-(aq) + \text{Fe}^{3+}(aq) \rightarrow \text{Fe(OH)}_3(s) \]

See Exercises 4.47 through 4.52

---

### 4.7 Stoichiometry of Precipitation Reactions

In Chapter 3 we covered the principles of chemical stoichiometry: the procedures for calculating quantities of reactants and products involved in a chemical reaction. Recall that in performing these calculations we first convert all quantities to moles and then use the coefficients of the balanced equation to assemble the appropriate mole ratios. In cases where reactants are mixed, we must determine which reactant is limiting, since the reactant that is consumed first will limit the amounts of products formed. These same principles apply to reactions that take place in solutions. However, two points about solution reactions need special emphasis. The first is that it is sometimes difficult to tell immediately what reaction will occur when two solutions are mixed. Usually we must do some thinking about the various possibilities and then decide what probably will happen. The first step in this process always should be to write down the species that are actually present in the solution, as we did in Section 4.5. The second special point about solution reactions is that to obtain the moles of reactants we must use the volume of the solution and its molarity. This procedure was covered in Section 4.3.

We will introduce stoichiometric calculations for reactions in solution in Example 4.10.

**Critical Thinking**

What if all ionic solids were soluble in water? How would this affect stoichiometry calculations for reactions in aqueous solution?

---

### Determining the Mass of Product Formed

Calculate the mass of solid NaCl that must be added to 1.50 L of a 0.100-M AgNO₃ solution to precipitate all the Ag⁺ ions in the form of AgCl.
Predicting Reaction Products

Using the solubility rules in Table 4.1, predict what will happen when the following pairs of solutions are mixed.

a. $\text{KNO}_3(aq)$ and $\text{BaCl}_2(aq)$

b. $\text{Na}_2\text{SO}_4(aq)$ and $\text{Pb(NO}_3\text{)}_2(aq)$

c. $\text{KOH}(aq)$ and $\text{Fe(NO}_3\text{)}_3(aq)$

Solution

a. The formula $\text{KNO}_3(aq)$ represents an aqueous solution obtained by dissolving solid $\text{KNO}_3$ in water to form a solution containing the hydrated ions $\text{K}^+(aq)$ and $\text{NO}_3^-(aq)$. Likewise, $\text{BaCl}_2(aq)$ represents a solution formed by dissolving solid $\text{BaCl}_2$ in water to produce $\text{Ba}^{2+}(aq)$ and $\text{Cl}^-(aq)$. When these two solutions are mixed, the resulting solution contains the ions $\text{K}^+$, $\text{NO}_3^-$, $\text{Ba}^{2+}$, and $\text{Cl}^-$. All ions are hydrated, but the $(aq)$ is omitted for simplicity. To look for possible solid products, combine the cation from one reactant with the anion from the other:

$$\text{K}^+ + \text{NO}_3^- + \text{Ba}^{2+} + \text{Cl}^- \rightarrow$$

Note from Table 4.1 that the rules predict that both $\text{KCl}$ and $\text{Ba(NO}_3\text{)}_2$ are soluble in water. Thus no precipitate forms when $\text{KNO}_3(aq)$ and $\text{BaCl}_2(aq)$ are mixed. All the ions remain dissolved in solution. No chemical reaction occurs.

b. Using the same procedures as in part a, we find that the ions present in the combined solution before any reaction occurs are $\text{Na}^+$, $\text{SO}_4^{2-}$, $\text{Pb}^{2+}$, and $\text{NO}_3^-$. The possible salts that could form precipitates are

$$\text{Na}^+ + \text{SO}_4^{2-} + \text{Pb}^{2+} + \text{NO}_3^- \rightarrow$$

The compound $\text{NaNO}_3$ is soluble, but $\text{PbSO}_4$ is insoluble (see Rule 4 in Table 4.1). When these solutions are mixed, $\text{PbSO}_4$ will precipitate from the solution. The balanced equation is

$$\text{Na}_2\text{SO}_4(aq) + \text{Pb(NO}_3\text{)}_2(aq) \rightarrow \text{PbSO}_4(s) + 2\text{NaNO}_3(aq)$$

c. The combined solution (before any reaction occurs) contains the ions $\text{K}^+$, $\text{OH}^-$, $\text{Fe}^{3+}$, and $\text{NO}_3^-$. The salts that might precipitate are $\text{KNO}_3$ and $\text{Fe(OH)}_3$. The solubility rules in Table 4.1 indicate that both $\text{K}^+$ and $\text{NO}_3^-$ salts are soluble. However, $\text{Fe(OH)}_3$ is only slightly soluble (Rule 5) and hence will precipitate. The balanced equation is

$$3\text{KOH}(aq) + \text{Fe(NO}_3\text{)}_3(aq) \rightarrow \text{Fe(OH)}_3(s) + 3\text{KNO}_3(aq)$$

See Exercises 4.45 and 4.46

4.6 | Describing Reactions in Solution

In this section we will consider the types of equations used to represent reactions in solution. For example, when we mix aqueous potassium chromate with aqueous barium nitrate, a reaction occurs to form a precipitate ($\text{BaCrO}_4$) and dissolved potassium nitrate. So far we have written the overall or formula equation for this reaction:

$$\text{K}_2\text{CrO}_4(aq) + \text{Ba(NO}_3\text{)}_2(aq) \rightarrow \text{BaCrO}_4(s) + 2\text{KNO}_3(aq)$$
A dilution procedure typically involves two types of glassware: a pipet and a volumetric flask. A pipet is a device for accurately measuring and transferring a given volume of solution. There are two common types of pipets: volumetric (or transfer) pipets and measuring pipets (Fig. 4.11). Volumetric pipets come in specific sizes, such as 5 mL, 10 mL, 25 mL, and so on. Measuring pipets are used to measure volumes for which a volumetric pipet is not available. For example, we would use a measuring pipet as shown in Fig. 4.12 on page 153 to deliver 28.7 mL of 17.4 M acetic acid into a 500-mL volumetric flask and then add water to the mark to perform the dilution described above.

Concentration and Volume

What volume of 16 M sulfuric acid must be used to prepare 1.5 L of a 0.10-M H₂SO₄ solution?

Solution

Where are we going?

To find the volume of H₂SO₄ required to prepare the solution

What do we know?

- 1.5 L of 0.10 M H₂SO₄ is required
- We have 16 M H₂SO₄

What information do we need to find the volume of H₂SO₄?

- Moles of H₂SO₄ in the required solution

How do we get there?

What are the moles of H₂SO₄ required?

\[ M \times V = \text{mol} \]
\[ 1.5 \text{ L solution} \times \frac{0.10 \text{ mol H}_2\text{SO}_4}{1 \text{ L solution}} = 0.15 \text{ mol H}_2\text{SO}_4 \]

What volume of 16 M H₂SO₄ contains 0.15 mole of H₂SO₄?

\[ V \times \frac{16 \text{ mol H}_2\text{SO}_4}{1 \text{ L solution}} = 0.15 \text{ mol H}_2\text{SO}_4 \]

Solving for \( V \) gives

\[ V = \frac{0.15 \text{ mol H}_2\text{SO}_4}{16 \text{ mol H}_2\text{SO}_4} = 9.4 \times 10^{-3} \text{ L or } 9.4 \text{ mL solution} \]

To make 1.5 L of 0.10 M H₂SO₄ using 16 M H₂SO₄, we must take 9.4 mL of the concentrated acid and dilute it with water to 1.5 L. The correct way to do this is to add the 9.4 mL of acid to about 1 L of distilled water and then dilute to 1.5 L by adding more water.

See Exercises 4.35b,d, and 4.36a,b,d

As noted earlier, the central idea in performing the calculations associated with dilutions is to recognize that the moles of solute are not changed by the dilution. Another way to express this condition is by the following equation:

\[ M_1V_1 = M_2V_2 \]
What do we know?
> 1.00 L of 0.200 M $K_2Cr_2O_7$ is required

What information do we need to find the mass of $K_2Cr_2O_7$?
> Moles of $K_2Cr_2O_7$ in the required solution

How do we get there?
What are the moles of $K_2Cr_2O_7$ required?

\[ M \times V = \text{mol} \]
\[ 1.00 \text{ L solution} \times \frac{0.200 \text{ mol } K_2Cr_2O_7}{1 \text{ L solution}} = 0.200 \text{ mol } K_2Cr_2O_7 \]

What mass of $K_2Cr_2O_7$ is required for the solution?

\[ 0.200 \text{ mol } K_2Cr_2O_7 \times \frac{294.20 \text{ g } K_2Cr_2O_7}{1 \text{ mol } K_2Cr_2O_7} = 58.8 \text{ g } K_2Cr_2O_7 \]

To make 1.00 L of 0.200 M $K_2Cr_2O_7$, the chemist must weigh out 58.8 g $K_2Cr_2O_7$, transfer it to a 1.00-L volumetric flask, and add distilled water to the mark on the flask.

**Dilution**

To save time and space in the laboratory, routinely used solutions are often purchased or prepared in concentrated form (called stock solutions). Water is then added to achieve the molarity desired for a particular solution. This process is called dilution. For example, the common acids are purchased as concentrated solutions and diluted as needed. A typical dilution calculation involves determining how much water must be added to an amount of stock solution to achieve a solution of the desired concentration. The key to doing these calculations is to remember that

Moles of solute after dilution = moles of solute before dilution

because only water (no solute) is added to accomplish the dilution.

For example, suppose we need to prepare 500. mL of 1.00 M acetic acid ($HC_2H_3O_2$) from a 17.4-M stock solution of acetic acid. What volume of the stock solution is required? The first step is to determine the number of moles of acetic acid in the final solution by multiplying the volume by the molarity (remembering that the volume must be changed to liters):

\[ \frac{500. \text{ mL solution}}{1000 \text{ mL solution}} \times \frac{1 \text{ L solution}}{1000 \text{ mL solution}} \times \frac{1.00 \text{ mol } HC_2H_3O_2}{1 \text{ L solution}} = 0.500 \text{ mol } HC_2H_3O_2 \]

Thus we need to use a volume of 17.4 M acetic acid that contains 0.500 mole of $HC_2H_3O_2$. That is,

\[ V \times \frac{17.4 \text{ mol } HC_2H_3O_2}{1 \text{ L solution}} = 0.500 \text{ mol } HC_2H_3O_2 \]

Solving for $V$ gives

\[ V = \frac{0.500 \text{ mol } HC_2H_3O_2}{17.4 \text{ mol } HC_2H_3O_2} \times \frac{1.00 \text{ L solution}}{0.0287 \text{ L or 28.7 mL solution}} = 0.0287 \text{ L or 28.7 mL solution} \]

Thus to make 500 mL of a 1.00-M acetic acid solution, we can take 28.7 mL of 17.4 M acetic acid and dilute it to a total volume of 500 mL with distilled water.
4.3 The Composition of Solutions

What information do we need to find volume of blood containing 1.0 mg of NaCl?

- Moles of NaCl (in 1.0 mg)

How do we get there?

What are the moles of NaCl (58.44 g/mol)?

\[
1.0 \text{ mg NaCl} \times \frac{1 \text{ g NaCl}}{1000 \text{ mg NaCl}} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 1.7 \times 10^{-5} \text{ mol NaCl}
\]

What volume of 0.14 M NaCl contains 1.0 mg (1.7 \times 10^{-5} mole) of NaCl?

There is some volume, call it V, that when multiplied by the molarity of this solution will yield 1.7 \times 10^{-5} mole of NaCl. That is,

\[V \times \frac{0.14 \text{ mol NaCl}}{\text{L solution}} = 1.7 \times 10^{-5} \text{ mol NaCl}\]

We want to solve for the volume:

\[V = \frac{1.7 \times 10^{-5} \text{ mol NaCl}}{0.14 \text{ mol NaCl}} = 1.2 \times 10^{-4} \text{ L solution}\]

Thus 0.12 mL of blood contains 1.7 \times 10^{-5} mole of NaCl or 1.0 mg of NaCl.

See Exercises 4.33 and 4.34

A standard solution is a solution whose concentration is accurately known. Standard solutions, often used in chemical analysis, can be prepared as shown in Fig. 4.10 and in Example 4.6.

Interactive Example 4.6

Sign in at http://login.cengagebrain.com to try this Interactive Example in OWL.

Solution

Where are we going?

To find the mass of K_2Cr_2O_7 required for the solution

Figure 4.10 | Steps involved in the preparation of a standard aqueous solution. (a) Put a weighed amount of a substance (the solute) into the volumetric flask, and add a small quantity of water. (b) Dissolve the solid in the water by gently swirling the flask (with the stopper in place). (c) Add more water (with gentle swirling) until the level of the solution just reaches the mark etched on the neck of the flask. Then mix the solution thoroughly by inverting the flask several times.

A standard solution is a solution whose concentration is accurately known. Standard solutions, often used in chemical analysis, can be prepared as shown in Fig. 4.10 and in Example 4.6.

See Exercises 4.33 and 4.34
**Interactive Example 4.3**

Sign in at [http://login.cengagebrain.com](http://login.cengagebrain.com) to try this Interactive Example in OWL.

**Concentration of Ions**

Give the concentration of each type of ion in the following solutions:

- a. 0.50 \( M \) \( \text{Co(NO}_3\text{)}_2 \)
- b. 1 \( M \) \( \text{Fe(ClO}_4\text{)}_3 \)

**Solution**

**Where are we going?**

To find the molarity of each ion in the solution

**What do we know?**

- 0.50 \( M \) \( \text{Co(NO}_3\text{)}_2 \)
- 1 \( M \) \( \text{Fe(ClO}_4\text{)}_3 \)

**What information do we need to find the molarity of each ion?**

- Moles of each ion

**How do we get there?**

For \( \text{Co(NO}_3\text{)}_2 \)

What is the balanced equation for dissolving the ions?

\[ \text{Co(NO}_3\text{)}_2(s) \xrightarrow{\text{H}_2\text{O}} \text{Co}^{2+}(aq) + 2\text{NO}_3^-(aq) \]

**What is the molarity for each ion?**

- \( \text{Co}^{2+} \) \( 1 \times 0.50 \, M = 0.50 \, M \) \( \text{Co}^{2+} \)
- \( \text{NO}_3^- \) \( 2 \times 0.50 \, M = 1.0 \, M \) \( \text{NO}_3^- \)

For \( \text{Fe(ClO}_4\text{)}_3 \)

What is the balanced equation for dissolving the ions?

\[ \text{Fe(ClO}_4\text{)}_3(s) \xrightarrow{\text{H}_2\text{O}} \text{Fe}^{3+}(aq) + 3\text{ClO}_4^- (aq) \]

**What is the molarity for each ion?**

- \( \text{Fe}^{3+} \) \( 1 \times 1 \, M = 1 \, M \) \( \text{Fe}^{3+} \)
- \( \text{ClO}_4^- \) \( 3 \times 1 \, M = 3 \, M \) \( \text{ClO}_4^- \)