Look at the values given for KCl solutions in Table 14-3. The freezing-point depression of a 0.1 m KCl solution is only 1.85 times greater than that of a nonelectrolyte solution. However, as the concentration decreases, the freezing-point depression comes closer to the value that is twice that of a nonelectrolyte solution.

The differences between the expected and calculated values are caused by the attractive forces that exist between dissociated ions in aqueous solution. The attraction between the hydrated ions in the solution is small compared with those in the crystalline solid. However, forces of attraction do interfere with the movements of the aqueous ions. The more concentrated a solution is, the closer together the ions are, and the greater the attraction between ions is. Only in very dilute solutions is the average distance between the ions large enough and the attraction between ions small enough for the solute ions to move about almost completely freely.

Peter Debye and Erich Hückel introduced a theory in 1923 to account for this attraction between ions in dilute aqueous solutions. According to this theory, the attraction between dissociated ions of ionic solids in dilute aqueous solutions is caused by an ionic atmosphere that surrounds each ion. This means that each ion, on average, is surrounded by more ions of opposite charge than of like charge. This clustering effect hinders the movements of solute ions. A cluster of hydrated ions can act as a single unit rather than as individual ions. Thus, the effective total concentration is less than expected, based on the number of ions known to be present. Ions of higher charge attract other ions very strongly. They therefore cluster more and have lower effective concentrations than ions with smaller charge. For example, ions formed by MgSO4 have charges of 2+ and 2−. Ions formed by KCl have charges of 1+ and 1−. Note in Table 14-3 that MgSO4 in a solution does not depress the freezing point as much as the same concentration of KCl.

**SECTION REVIEW**

1. What colligative properties are displayed by each of the following situations?
   a. Antifreeze is added to a car’s cooling system to prevent freezing when the air temperature is below 0°C.
   b. Ice melts on sidewalks after salt has been spread on them.

2. Two moles of a nonelectrolytic solute are dissolved in 1 kg of an unknown solvent. The solution freezes at 7.8°C below its normal freezing point. What is the molar freezing-point constant of the unknown solvent? What is your prediction for the identity of the solvent?

3. If two solutions of equal amounts in a U-tube are separated by a semipermeable membrane, will the level of the more-concentrated solution or the less-concentrated solution rise?

4. a. Calculate the expected freezing-point depression of a 0.2 m KNO3 solution.
b. Will the value you calculated match the actual freezing-point depression for this solution? Why or why not?

**CHAPTER SUMMARY**

14-1

- The separation of ions that occurs when an ionic solid dissolves is called dissociation.
  + When two different ionic solutions are mixed, a precipitate may form if ions from the two solutions react to form an insoluble compound.
  + A net ionic equation for a reaction in aqueous solution includes only compounds and ions that change chemically in the reaction. Spectator ions are ions that do not take part in such a reaction.

**Vocabulary**

- dissociation (425)
- ionization (431)
- hydrion (431)
- net ionic equation (429)

14-2

- Colligative properties of solutions depend only on the total number of solute particles present.
  - Boiling-point elevation, freezing-point depression, vapor-pressure lowering, and osmotic pressure are all colligative properties.
  - The molal boiling-point and freezing-point constants are used to calculate boiling-point elevations and freezing-point depressions of solutions containing nonvolatile solutes.

**Vocabulary**

- boiling-point elevation, Δb (446)
- colligative properties (436)
- freezing-point depression, Δf (438)
- molal boiling-point constant, Kb (440)
- molal freezing-point constant, Kf (438)

- Formation of ions from solute molecules is called ionization. A molecular compound may ionize in a water solution if the attraction of the polar water molecules is strong enough to break the polar-covalent bonds of the solute molecules.
  - An H2O+ ion is called a hydronium ion.
  - All, or almost all, of a dissolved strong electrolyte exists as ions in an aqueous solution, whereas a relatively small amount of a dissolved weak electrolyte exists as ions in an aqueous solution.

**Reviewing Concepts**

1. How many moles of ions are contained in 1 L of a 1 M solution of KCl? of Mg(NO3)2?

2. Use Table 14-1 to predict whether each of the following compounds is considered soluble or insoluble:
   a. KCl
   b. NaNO3
   c. AgCl

- BaSO4
- Ca(H2PO4)2
- Pb(NO3)2
- (NH4)2S
- PbCl2 (in cold water)
- FeS
- Al2(SO4)3

3. What is a net ionic equation?

4. a. What is ionization?
b. Distinguish between ionization and dissociation.
5. a. Define and distinguish between strong electrolytes and weak electrolytes.
   b. Give two examples of each type.

6. What determines the strength with which a solute acts as an electrolyte?

7. Distinguish between the use of the terms strong and weak and the use of the terms dilute and concentrated when used to describe electrolyte solutions.

8. How does the presence of a nonvolatile solute affect each of the following properties of the solvent into which the solute is dissolved?
   a. vapor pressure
   b. freezing point
   c. boiling point
   d. osmotic pressure

9. Using Figure 14-6 as a guide, make a sketch of a vapor pressure-versus-temperature curve that shows the comparison of pure water, a solution with x amount of solute, and a solution with 2x the amount of solute. What is the relationship between ΔG for the x curve and ΔG for the 2x curve?

10. a. Why does the level of the more-concentrated solution rise when two solutions of different concentrations are separated by a semipermeable membrane?
    b. When does the level of the solution stop rising?
    c. When the level stops rising, what is the net movement of water molecules across the membrane?

11. a. Compare the effects of nonvolatile electrolytes with the effects of nonvolatile nonelectrolytes on the freezing and boiling points of solvents in which they are dissolved.
    b. Why are such differences observed?

12. Why does the actual freezing-point depression of an electrolytic solution differ from the freezing-point depression calculated on the basis of the concentration of particles?

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**Problems**

### Dissociation

13. Write the equation for the dissociation of each of the following ionic compounds in water. (Hint: See Sample Problem 14-1.)
   a. KI
   b. NaNO₃
   c. MgCl₂
   d. Na₂SO₄

14. For the compounds listed in the previous problem, determine the number of moles of each ion produced as well as the total number of moles of ions produced when 1 mol of each compound dissolves in water.

15. Write the equation for the dissociation of each of the following in water, and then indicate the total number of moles of solute ions formed.
   b. 0.50 mol strontium nitrate
   c. 0.50 mol sodium phosphate

### Precipitation Reactions

16. Using Table 14-1, write the balanced chemical equation, write the overall ionic equation, identify the spectator ions and possible precipitates, and write the net ionic equation for each of the following reactions. (Hint: See Sample Problem 14-2.)
   a. mercury(II) chloride (aq) + potassium sulfide (aq) →
   b. sodium carbonate (aq) + calcium chloride (aq) →
   c. copper(II) chloride (aq) + ammonium phosphate (aq) →

17. Identify the spectator ions in the reaction between KCJ and AgNO₃ in an aqueous solution.

18. Copper(II) chloride and lead(II) nitrate react in aqueous solutions by double replacement. Write the balanced chemical equation, the overall ionic equation, and the net ionic equation for this reaction. If 13.45 g of copper(II) chloride react, what is the maximum amount of precipitate that could be formed?

### Freezing-Point Depression of Nonelectrolytes

19. Determine the freezing-point depression of H₂O in each of the following solutions. (Hint: See Sample Problem 14-3.)
   a. 150 m solution of C₂H₃O₂⁻ (sucrose) in H₂O
   b. 171 g of C₂H₅O₂⁻ in 1.00 kg H₂O
   c. 77.0 g of C₂H₅O₂⁻ in 400 g H₂O

20. Determine the molarity of each solution of an unknown nonelectrolyte in water, given the following freezing-point depressions. (Hint: See Sample Problem 14-4.)
   a. -0.90°C
   b. -3.72°C
   c. -8.37°C

21. A solution contains 20.0 g of C₂H₅O₂⁻ (glucose) in 250 g of water.
   a. What is the freezing-point depression of the solvent?
   b. What is the freezing point of the solution?

22. How many grams of antifreeze, C₄H₈O₂(OH)₄, would be required per 500 g of water to prevent the water from freezing at a temperature of -20.0°C?

23. Pure benzene, C₆H₆, freezes at 5.45°C. A solution containing 7.24 g of C₆H₆ in 115 g of benzene (specific gravity = 0.879) freezes at 3.55°C. Based on these data, what is the molar freezing-point constant for benzene?

24. If 1.50 g of a solute having a molar mass of 125.0 g were dissolved in 35.00 g of camphor, what would be the resulting freezing point of the solution?

### Boiling-Point Elevation of Nonelectrolytes

25. Determine the boiling-point elevation of H₂O in each of the following solutions. (Hint: See Sample Problem 14-5.)
   a. 2.5 m solution of C₉H₁₈O₄⁻ (glucose) in H₂O
   b. 3.20 g of C₉H₁₈O₄⁻ in 1.00 kg H₂O
   c. 20.0 g of C₉H₁₈O₄⁻ (sucrose) in 500 g H₂O

26. Determine the molality of each water solution given the following boiling points:
   a. 100.25°C
   b. 101.53°C
   c. 102.80°C

### Colligative Properties of Electrolytes

27. Given 1.00 m aqueous solutions of each of the following electrolytic substances, what is the expected change in the freezing point of the solvent? (Hint: See Sample Problem 14-6.)
   a. KI
   b. CaCl₂
   c. Ba(NO₃)₂

28. What is the expected change in the freezing point of water for an aqueous solution that is 0.015 m AICl₃?

29. What is the expected freezing point of a solution containing 85.0 g of NaCl dissolved in 450 g of water?

30. Determine the expected boiling point of a solution made by dissolving 25.0 g of barium chloride in 0.150 kg of water.

31. The change in the boiling point of water for an aqueous solution of potassium iodide is 0.65°C. Determine the apparent molal concentration of potassium iodide.

32. The freezing point of an aqueous solution of barium nitrate is -2.65°C. Determine the apparent molal concentration of barium nitrate.

33. Calculate the expected freezing point of a solution containing 1.00 kg of H₂O and 0.250 mol of NaCl.

34. Experimental data for a 1.00 m Mg₃PO₄ aqueous solution indicate an actual change in the freezing point of water of -4.76°C. Determine the expected change in the freezing point of water. Suggest a possible reason for discrepancies between the experimental and the expected values.

### Mixed Review

35. Given 0.01 m aqueous solutions of each of the following, arrange the solutions in order of increasing change in the freezing point of the solvent.
   a. NaCl
   b. CaCl₂
   c. K₂PO₄
   d. C₂H₅O₂⁻ (glucose)
Critical Thinking

46. Applying Models
   a. You are conducting a freezing-point determination in the laboratory using an aqueous solution of KNO₃. The observed freezing point of the solution is -1.15°C. Using a pure water sample, you recorded the freezing point of the pure solvent on the same thermometer as 0.15°C. Determine the molal concentration of KNO₃. Assume that there are no forces of attraction between ions.
   b. You are not satisfied with the result in part (a) because you suspect that you should not ignore the effect of ion interaction. You take a 10.00 mL sample of the solution. After carefully evaporating the water from the solution, you obtain a mass of 0.415 g KNO₃. Determine the actual molal concentration of KNO₃ and the percentage difference between the predicted concentration and the actual concentration of KNO₃. Assume that the solution's density is 1.00 g/mL.

47. Analyzing Information
   a. Nitric acid, HNO₃, is a weak electrolyte. Nitric acid, HNO₂, is a strong electrolyte. Write equations to represent the ionization of each in water. Include the hydronium ion, and show the appropriate kind of arrow in each case.
   b. Find the boiling point of an aqueous solution containing a nonelectrolyte that freezes at -6.5°C.

48. Analyzing Information
   a. Write a balanced equation for the dissolution of sodium carbonate, Na₂CO₃, in water. Find the number of moles of each ion produced when 0.20 mol of sodium carbonate dissolves. Then find the total number of moles of ions.
   b. Given the reaction below and the information in Table 14-1, write the net ionic equation for the reaction.

49. Common reactions for Group 13 elements are found in the Elements Handbook. Review this material and answer the following.
   a. Write net ionic equations for each of the example reactions shown on page 751.
   b. Which reactions did not change when written in net ionic form? Why?

50. Common reactions for Group 14 elements are found in the Elements Handbook. Review this material and answer the following.
   a. Write net ionic equations for each of the example reactions shown on page 755.
   b. Which reactions did not change when written in net ionic form? Why?

Research & Writing

51. Find out how much salt a large northern city, such as New York City or Chicago, uses on its streets in a typical winter. What environmental problems result from this use of salt? What substitutes for salt are being used to melt ice and snow?

52. Research the role of electrolytes and electrolyte solutions in your body. Find out how electrolytes work in the functioning of nerves and muscles. What are some of the health problems that can arise from an imbalance of electrolytes in body fluids?