Answer Key

Unit 4 Solutions and Solubility

Unit Preparation Questions (Assessing Readiness)

(Student textbook pages 348–51)

1. a. eye hazard; safety goggles or a face shield should be worn

b. corrosive substances, such as acids and bases, that can cause burns or are poisonous if absorbed through the skin

c. protective gloves need to be worn

2. a. Material Safety Data Sheet and Workplace Hazardous Materials Information System

b. A chemical's MSDS lists all of the chemical's physical properties (such as melting point and boiling point), health effects, storage, and disposal, and information on how to safely handle the chemical.

c. *WHMIS* is a Canadian system that allows people to get information—including safety information—about hazardous materials used in their workplace.

d. Sample answer: An alphabetical file of MSDSs should be located at the front of the classroom.

3. Sample answer: I would first tell my lab partner to hold his arm under running water to wash off the chemical. I would then tell my teacher what happened. Next, I would help clean up the spill according to my teacher's directions. I would also clean up the broken glassware and place the pieces in the broken glassware bin.

4. a. mixture **b.** element **c.** compound **d.** mixture **e.** element **f.** compound

5. d

6. Compounds can be broken down into simpler substances (i.e., elements) by chemical reactions. Mixtures can be broken down into their components through physical processes such as filtering or dissolving.

7. Sample answer: Examples of heterogeneous mixtures are pizza, soil, and rocks. Examples of homogenous mixtures are air, lemonade, and seawater.

8. a. Greater, because the higher edge is being viewed, rather than the bottom of the meniscus.

b. The eye should be lowered to the same level as the bottom of the meniscus.

c. 18.5 ± 0.1 mL

9. a. 0.0268 L **b.** 355 mL **c.** 1.25 dL **d.** 100 mL

10. The first way to measure the volume of the marble is to measure its diameter using a ruler (or a set of calipers) and the divide it by 2 to find the radius. Then use the radius and the formula $V = 4/3\pi r^3$ to find the volume. The other way to find the volume is to use water displacement. Partly fill a graduated cylinder with water, read the initial volume, add the marble, and read the final volume. Subtract the initial volume from the final volume to find the volume of the marble.

11. e

12. A decomposition reaction occurred. Water is being broken down into hydrogen and oxygen gases. These gases are being collected inside the test tubes.

13. A synthesis reaction is like two dance partners getting together to dance. A decomposition reaction is like two dance partners separating after a dance is over. A single displacement reaction is like one person cutting in on a pair of dance partners and taking the place of one of the partners. A double displacement reaction is like two sets of dance partners switching partners.

14. d

15. d

16. c

17. b

18. a. lead(II) iodide, $PbI_2(s)$

b. double displacement

c. The limiting reactant is KI(aq) and 0.15 mol $PbI_2(s)$ will precipitate.

19. d

20. c

21. a. $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$; acid

b. $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$; base

c. $Ca(OH)_2(aq) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$; base

d. $HClO_3(aq) \rightarrow H^+(aq) + ClO_3^-(aq)$; acid

22. The pH of an acidic solution is less than 7. As a base is added, the pH will increase. The pH of the solution may become greater than 7 if an excess of base is added.

23. drain cleaner, soap; Bases turn red litmus paper blue.

24. c

25. b

26. **a.** acid
    b. base
    c. base
    d. both
    e. base
    f. acid
    g. both
    h. both
    i. acid

27. The pH is 7 because equal amounts of acid and base reacted in a neutralization reaction.

Chapter 8 Solutions and Their Properties

Learning Check Questions

(Student textbook page 357)

1. The mixtures are not homogeneous on the microscopic scale.

2. Nitrogen is the solvent, because it is present in the largest proportion.

3. Sample answer: solid solute: sugar in water; liquid solute: alcohol in water; gas solute: carbon dioxide in water.

4. Gasoline and water do not dissolve in one another (i.e., they are immiscible).

5. Yes, a solute is classified as insoluble if less than 0.1 g will dissolve in 100 mL of solvent. Therefore, if as much solute is dissolved in the solvent as possible, the solution will be dilute but saturated.

6. Yes, when the pressure is released by opening the can, the excess solute (carbon dioxide gas) leaves the solution.

(Student textbook page 368)

7. No, like dissolves like, so if benzene is a solvent for fats and oils, which are non-polar, it should not dissolve in a polar solvent such as water.

8. A fluoride ion has a very small radius. This results in most fluorides being less soluble than the other halides.

9. Methanol. The solubility of liquids is changed very little by temperature.

10. The taste is affected by how much carbon dioxide is dissolved in the soft drink. The solubility of carbon dioxide in water decreases at a higher temperature.

11. Unsaturated. At 40°C, the solubility of $KNO_3(s)$ is about 53 g/100 g.

12. Fish avoid extremes of hot and cold, just as we do. In particular, warm water contains less dissolved oxygen, so fish are less likely to be found in these areas.

(Student textbook page 388)

13. The concentration is likely to be more precise.

14. volumetric flask and volumetric pipette, or graduated pipette

15. (1) Measure the mass of pure solute and dissolve in water. Transfer to a 1 L volumetric flask and make up the solution to the calibration mark.
    (2) Dilute a more concentrated standard solution using a graduated pipette and a 1 L volumetric flask.

16. $n = c \times V$; the product of concentration and volume is constant. Thus, as the volume of solution increases, its concentration decreases.

17. Drying the beaker is not necessary in this case. The water in the beaker will dilute the solution slightly, but it will not change the amount of $CuSO_4(aq)$, which is determined by the volume of the pipette and the concentration of the solution in it.

18. First prepare a more concentrated solution and dilute it to the necessary concentration.

Caption Questions

Figure 8.8 (Student textbook page 361): Sucrose is not an ionic compound.

Figure 8.11 (Student textbook page 365): The solubility is about 45 g KCl(aq) per 100 g water.

Figure 8.17 (Student textbook page 374): Iron is present in the greatest amount in the solution.

Figure 8.19 (Student textbook page 379): Sample answer: A pair of glasses has a frame and two lenses.

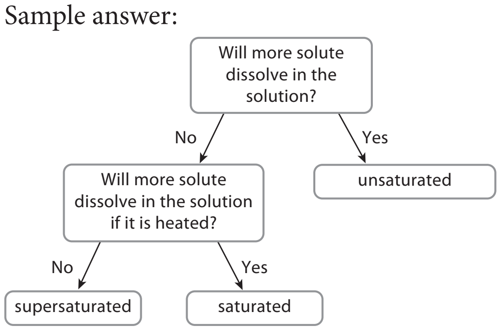
Figure 8.20 (Student textbook page 383): All of the flasks have long, thin necks, stoppers, and a graduation mark to indicate volume.

Section 8.1 Review Questions

(Student textbook page 358)

1. **a.** seawater
   b. The solvent is the component present in the greatest amount, so there can be only one.

2. Stainless steel, seawater, and air are mixtures of substances and are homogenous at the microscopic level.
3. No, the mixture is not homogenous at the microscopic level.
4. A solvent is the substance in a solution that is present in the greatest amount. Any other substance is a solute.
5. A solute is insoluble if less than 0.1 g of a solute will dissolve in 100 mL of solvent (e.g., oxygen in water). A solution is unsaturated if more of the same solute in the solution can be added to the solution (e.g., an unsaturated solution of sodium chloride in water. More sodium chloride can be added before the solution becomes saturated.).
6. **a.** colouring dissolved in water, carbon tetrachloride dissolved in wax
   b. If the solution were miscible the wax would distribute evenly throughout the container, making a uniform solution with no moving blobs.
7. Distilled water; water and ethanol; oxygen dissolved in water; aqueous sugar solution.
8. Heat each solution and measure the boiling point. Water will boil at 100°C, and the solution will boil at a different temperature. Alternatively, measure the density of each liquid. Water has a density of 1.00 g/mL, greater than the density of the solution.
9. The honey may be stored at a low temperature; crystals may form around crumbs of toast; if the lid is not airtight, some water may evaporate.
10. Cooling of the air and adding solute by evaporation of water on the ground both cause the air to become saturated.
11. Sample answer:



12. Because sodium chloride is readily soluble in water, the run-off of melting ice and snow will carry dissolved salt into soil, groundwater, and nearby bodies of water.
13. Sample answer: Try to dissolve a 10 g sample of each sample in 10 mL of water. The sucrose will dissolve completely, but less than half of the sodium chloride will dissolve.
14. The addition of a single crystal of solute added to a supersaturated liquid solution of the solute will result in the excess solute precipitating, which will leave the solution saturated.
15. Solute A is more soluble. For solute B, $(26/80) \times 100 = 32.5$. Thus, only 32.5 g of solute B dissolves in 100 mL of water.

Section 8.2 Review Questions

(Student textbook page 370)

1. Hydration involves a number of water molecules surrounding dissolved molecules of sugar. The polar water molecules are attracted to the polar –OH bonds on the sugar molecule.
2. **a.** Water is a polar compound. The molecular dipole attracts ions in an ionic substance. Covalent substances that contain the atoms F, O, or N can hydrogen bond to water molecules.
   b. No solvent dissolves all solutes, but water does dissolve a great variety of compounds.
3. The people do not have enough fat in their diet to dissolve the vitamin.
4. Both salts contain the OH^- ion, which is relatively small. The Ca^{2+} ion has twice the charge of the sodium ion. Increased charge and small ion size in $Ca(OH)_2$ result in a relatively strong ionic bond which is hard to break by water molecules, and the salt is insoluble.
5. In carbon tetrachloride, because both iodine and CCl_4 are non-polar.
6. Both olive oil and animal fat are non-polar. The olive oil can be used to estimate how much of the anaesthetic gas would dissolve in the fat of a person who breathes the gas.
7. Polar; they must dissolve in water.
8. The air pressure will be increased to match the pressure of air the diver was breathing at depth. This will allow the nitrogen to dissolve in the blood. Then, the air pressure will be decreased in steps over a period of time to allow the gas to leave the blood slowly.
9. **a.** supersaturated **b.** unsaturated **c.** saturated
10. **a.** soluble **b.** insoluble
11. Benzene is non-polar. The attraction between opposite ions in sodium chloride is much stronger than the attraction between the ions and benzene molecules.

12. Graphs should show that the solubility of a solid in water increases as the temperature increases (diagonal line, up and to the right), the solubility of a liquid in water is not affected by temperature (horizontal line), and the solubility of a gas in water decreases as the temperature increases (diagonally down and to the right).

13. The rate of dissolving refers to how quickly a given mass of solute dissolves in a solvent. Solubility refers to how much solute dissolves in a solvent at a given temperature. Temperature affects both the rate of dissolving and solubility.

14. **a.** temperature, agitation, particle size.
    b. temperature

15. The –OH part of the molecule is polar, and forms a hydrogen bond to a water molecule. The rest of the molecule is non-polar, enabling it to dissolve in non-polar solvents such as gasoline.

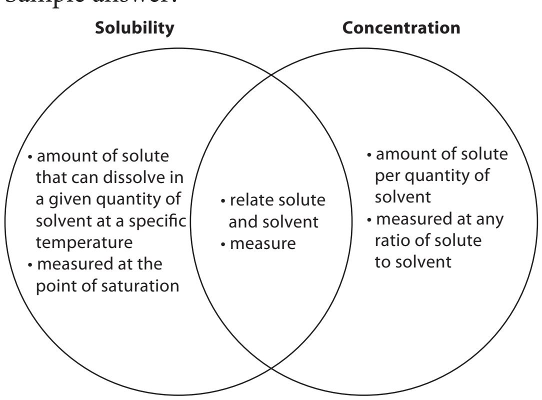
16. The solubility of most gases in water, including oxygen, decreases at higher temperatures. The boiled water will not contain enough dissolved oxygen for fish to live.

Section 8.3 Review Questions

(Student textbook page 382)

1. Diagrams should show an understanding that a concentrated solution has a much higher ratio of solute to solvent than a dilute solution does. Such as Figure 8.21 on page 385 of the student textbook.

2. Sample answer:



3. **a.** 0.04% (m/m) **b.** 4×10^2 ppm

4. 0.5 g

5. It is below the standard because it contains 0.4 ppb.

6. 23.9% (v/v)

7. 8.7 mol/L

8. The salt has dissociated into its component ions. The salt contains three potassium ions in every formula unit, e.g., K_3X. The salt could be potassium citrate, $K_3C_6H_5O_7$, or potassium phosphate, K_3PO_4.

9. 5.3×10^{14} kg

10. 0.7 mol/L

11. 0.098 mg

12. 45 ppb

13. No, the ground water concentration is over 300 ppb.

14. 0.4 mol/L

15. **a.** The phytoplankton are not dissolved in the seawater.
    b. Mass/volume concentration can be useful if the mixture is reasonably homogeneous. For example, if the phytoplankton are evenly distributed throughout the seawater.

Section 8.4 Review Questions

(Student textbook page 390)

1. A standard solution is one with accurately known concentration. A stock solution is a concentrated solution used to prepare more dilute solutions.

2. A volumetric flask is only accurate at room temperature. It would not accurately measure hot liquids.

3. 0.159 mol/L

4. Measure 33.3 mL of concentrated solution and dilute it with water until the final volume is 4.0 L. The procedure should be carried out in a fume hood. The teacher should wear gloves, goggles or a face shield, and a laboratory coat.

5. 3.00 L

6. 250 mL; 500 mL

7. 4.10 g

8. 7.5×10^9 L

9. **a.** The solution will be more dilute than expected. The rinse should be added to the volumetric flask.
   b. The solution will be more dilute than expected, because the meniscus appears to be higher than it really is. The eye should be level with the graduation mark.

10. Dilution may reduce the harm that could be caused by an acid, but it will increase run-off and spread contamination, make acid splash about, and generate a great deal of heat. Firefighter should protect themselves from splashes and from walking in the run-off. Breathing protection (BA) is also recommended as the acid may be vaporized.

11. The preparation must be done in a fume hood, wearing gloves, goggles or a face shield, and a laboratory coat. Measure 500 mL of acid. Pour the acid slowly into 1.5 L of water.

12. 567 mL

13. a. The solution made will be less concentrated than expected.
    b. The solution made will be less concentrated than expected.
    c. The solution made will be less concentrated than expected.

14. See the procedure in Table 8.7 on page 384 of the student textbook.

Quirks and Quarks: Antimony: The New Lead Questions

(Student textbook page 389)

1. Antimony is used as a flame retardant in many textiles and as a catalyst in many other consumer products made out of plastic. These products are sometimes incinerated, which releases particles of antimony compounds into the air. Antimony also leaches from plastic bottles.

2. Sample answer: Antimony is found in nature, in small quantities of the pure element (especially in Finland) and mixed in ore (from areas such as China, Russia, Bolivia, and South Africa).

   It is mined. It is also separated from copper, gold and silver ores by heating it (evaporating non-antimony components) or using a salt solute (to create antimony precipitate).

   Most antimony is used in flame retardants. It is also used in semiconductors, batteries (especially recyclable ones), paints, ceramic enamels, glass, pottery, and in anti-friction coatings. It has been used as an anti-parasitic, and as black eye make-up (in ancient Egypt).

   Antimony goes back into the environment as waste from manufacturing processes, and as it leaches (or is burned off) from garbage. Antimony can be recycled by the usual extraction methods described above, but it cannot be recycled in municipal recycling systems. Because of its value and rarity in nature, an industry has arisen which aims to recover antimony from scrap and waste (mostly batteries) in order to sell it.

3. Sample answer: A strong background in math, science, English, and geography can lead to a career in geochemistry as a mining geologist, environmental researcher, or nuclear geologist (especially in searching for isotope resources). University and/or college education in geology and/or chemistry are required.

   Often, advanced degrees and laboratory skills are required.

Practice Problems

(Student textbook page 373)

1. 1.2×10^2% (m/v)
2. 8.6% (m/v)
3. 31.5 g
4. 8.0×10^1 g
5. 6.67% (m/v)
6. 1.75% (m/v)
7. 131 g
8. 2.6 g of NaCl, 0.09 g of KCl, 0.1 g of $CaCl_2$
9. 800 mL
10. Measure 14 g of the solute and dissolve in water. Add water to bring the total volume of the solution to 400 mL.

(Student textbook page 375)

11. 15% (m/m)
12. 3.8% (m/m)
13. 90 kg chromium, 40 kg nickel, 370 kg iron
14. 7.91% (m/m)
15. 15.1% (m/m)
16. 11 g
17. 5 g
18. 1.7% (m/m)
19. 7.23×10^{-4}% (m/m)
20. 0.243% (m/m) nickel, 2.3% (m/m) copper, 2.3×10^{-4}% (m/m) platinum

(Student textbook page 376)

21. 3.5×10^2 mL
22. 16% (v/v)
23. 1.6 L
24. about 2.2 L
25. Add enough water to increase the volume to 950 mL
26. 1.5 L
27. 9.1% (v/v)
28. 6 L
29. 4.0×10^1% (v/v), assuming that the mixed solution has a volume of 20.0 L
30. 74 mL

(Student textbook page 378)

31. 7.2 ppm

32. 35 mg

33. 0.7 ppm

34. 2.5×10^{-6} g

35. 11.5 ppm

36. The concentration is 3.0 mg/L, within limits.

37. 0.125 g or 125 mg

38. 4.2 g

39. 0.52 μg or 5.2×10^{-7} g

40. 1×10^2 ppm or 100 ppm

(Student textbook page 381)

41. a. 1.5 mol/L **b.** 0.154 mol/L

42. 1.4 L

43. a. 1.2 mol/L **b.** 0.628 mol/L **c.** 0.096 mol/L

44. 0.14 g

45. a. 0.093 g **b.** 12 g **c.** 101 g

46. a. $Na^+(aq) = 1.2$ mol/L; $SO_4^{2-}(aq) = 0.60$ mol/L

b. $NH_4^+(aq) = 3.1$ mol/L; $PO_4^{3-}(aq) = 1.0$ mol/L

c. $Ca^{2+}(aq) = 1 \times 10^{-4}$ mol/L; $PO_4^{3-}(aq) =$ 7.6×10^{-5} mol/L

47. 0.29 mol/L

48. 12 mol/L

49. 218 mL

50. 2.0×10^{-5} mol/L

(Student textbook page 386)

51. a. 33.3 mL **b.** 107 mL **c.** 25 mL

52. a. 0.0994 mol/L **b.** 0.19 mol/L **c.** 0.0116 mol/L

53. 0.0938 L

54. 0.02 mol/L

55. a. 0.08 mol/L **b.** 0.17 mol/L

56. 15% (m/v)

57. 3.00×10^2 mL

58. 0.5 L; about 0.5 L

59. Procedures should be similar to the one outlined in Table 8.7 and use the following mass for each solid.

a. 2.1 g $AgNO_3(s)$

b. 6.05 g $K_2CO_3(s)$

c. 12.6 g $KMnO_4(s)$

60. Procedures should be similar to the one outlined in Table 8.8 and use the following amounts.

a. Add 29 mL of concentrate to 471 mL of water.

b. Add 8 mL of concentrate to 142 mL of water.

c. Add 945 mL of concentrate to 805 mL of water.

Chapter 8 Review Questions

(Student textbook pages 399–401)

1. c

2. b

3. d

4. d

5. a

6. b

7. d

8. c

9. turpentine

10. Oil and vinegar are immiscible. Shaking evens the distribution of particles throughout the mixture. Otherwise, the top layer would pour off first, not an even mixture.

11. Pressure has a negligible effect on the solubility of solids and liquids. The solubility of a gas increases with an increase in the pressure of that gas over water.

12. a. soluble

b. It is polar, as indicated by its solubility in water (“like dissolves like”).

13. a. soluble; All potassium salts are soluble.

b. insoluble; $CCl_4(\ell)$ is a non-polar substance.

c. soluble; All sodium salts are soluble.

14. Percentage concentrations are ratios of the mass or volume of solute to the mass or volume of solution. Molar concentrations indicate the amount of solute (in moles) dissolved in one litre of solution.

15. Volumetric pipette had a single graduation mark and is used to transfer a fixed volume of solution. Graduated pipette has graduations marks at regular intervals so it can be used to measure any volume within the range of these marks.

16. Protective gloves, goggles, lab coat. Do the dilution in a fume hood. Add acid to water

17. Ammonia is a polar molecule, thus liquid ammonia should dissolve ionic and polar substances.

18. a. 45 mL **b.** 64 mL

19. copper 95.1%, zinc 3.9%, tin 1.0%

20. 2.5 mg

21. 55.6 mol/L

22. a. 0.204% (m/v)

b. 860 ppm. You must assume the mass of 100 mL solution is 100 g.

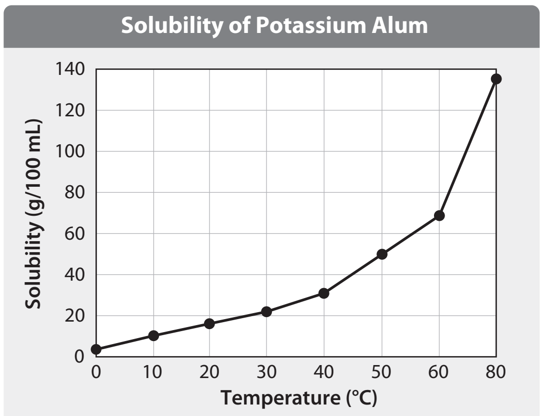
c. 0.14 mol/L

d. 1.0×10^{-3} mol/L

e. 4.5×10^{-3} mol/L

23. 2.2 L, assuming the volumes are additive

24. a.



b. approximately 90 g

c. approximately 76°C

25. a. approximately 380 g **b.** 40°C

26. Sample answer: To test the affect of surface area, test a number of samples of the same substance, as a single solid, cut into chunks, and powdered. Pour each sample into an equal amount of water and time how long it takes for each solid to disappear (dissolve). To test the affect of agitation, place identical solids into identical water samples. Let one sit undisturbed, stir one sample, and shake a third. Time how long each takes to dissolve. In all cases, keep the amount and temperature of water the same, as well as the size and type of container.

27. a. Diagrams should show that for the compound to be soluble, the attraction to the polar bonds in water must be stronger than the attraction between the ions in the compound. The shorter the distance between the ions is and the greater the charge on the ions, the stronger the bond between the ions will be. Thus small diameter ions and ions with greater charge are most likely to form insoluble compounds.

b. A covalent compound is likely to be soluble if it has polar bonds and the non-polar part of the molecule is not much larger than the polar part.

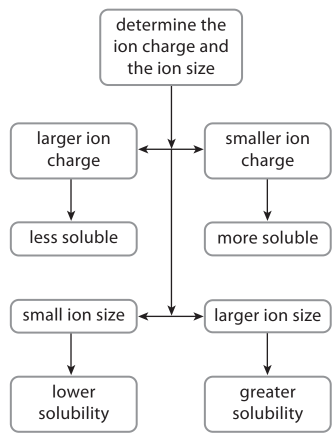
28. a. Diagrams should show that hydrogen bonds form between an oxygen atom on one water molecule and the hydrogen atoms of adjacent molecules as a result of dipole-dipole attractions. As in Figure 8.3 on page 359 of the student textbook.

b. The attraction between the dipoles on water molecules and the ions in a soluble ionic compound pull ions away from the surface of the ionic compound. The positive end of a dipole (hydrogen atom) attracts negative ions and the negative end of a dipole (oxygen atom) attracts positive ions.

The dipole attraction between polar covalent solute molecules is much weaker than hydrogen bonds between solute molecules and water molecules. Therefore, the water molecules easily separate the polar covalent molecules.

c. Non-polar compounds are only weakly attracted to water molecules so the attraction is too weak to break the hydrogen bonds between water molecules.

29. Sample answer:



30. Sample answers: Use a conductivity tester. Or, heat each liquid, and look for a solid residue in the solution containing the covalent substance. Or, cool both liquids and look for a precipitate in the solution.

31. Sample answer: The pond may support a lot of fish that have consumed oxygen in respiration, and few plants that are replacing the oxygen.

32. Diagrams should show an understanding that a solute will probably dissolve in a solvent with a similar polar (or non-polar) nature. They should show that the intermolecular forces between oil molecules, and between water molecules, is greater than the attraction between oil and water molecules. Thus, water cannot separate and hydrate oil molecules.
Polar molecules will tend to dissolve in polar solutes, because hydrogen bonding occurs in both. This means that hydrogen bonds can easily form between solute and solvent molecules. Hydrogen bonding does not occur between non-polar molecules. When a non-polar solute, such as oil, is placed in a polar solvent, such as water, hydrogen bonding between the polar solvent molecules will bind them together while the non-polar solute molecules are excluded. The same is true for polar solute molecules in a non-polar solvent.

33. Organizers should reflect the content of the chapter as summarized on page 398. Concepts should be organized from the most general to the most specific and include the key terms of the chapter. The organizer should show multiple levels (general to specific) and valid cross links among concepts, using appropriate linking words or symbols. The organizer should have an effective title, be easy to follow, and can show prior knowledge as well as new knowledge.

34. a. insoluble **b.** insoluble

c. Not a solution; the paint contains finely ground cinnabar in suspension; the mixture is not homogeneous at the microscopic scale.

35. a. 0.1 ppb **b.** 5×10^{-10} mol/L

c. The concentration is too small to be economically worthwhile. Huge volumes of water would have to be processed.

36. Add the mixture to water. Only solid A will dissolve. Filter the mixture and heat the solution to recover solid A. Add the filtered mixture to a non-polar solvent, perhaps mineral spirits. Only solid B will dissolve. Filter to recover solid C. Let the non-polar solvent evaporate and recover solid B. Note that many non-polar solvents are flammable, so heating to remove the solvent must be done with great care.

37. Since salt dissolves readily in water, the ground water that wicks up to the surface carries dissolved salt to the surface. At the surface, some of the water evaporates, leaving the salt as a surface deposit. If there is not enough rainwater to dissolve the salt and carry it back down into the soil, the salt deposit will gradually grow.

38. Milk is a colloidal emulsion. Raw milk is milk passed under high pressure through a tiny hole. This decreases the average size of the fat globules so they are similar in density to the rest of the milk.

39. a. radioactivity per volume, such as pCi/L (picocuries per litre) or Bq/m^3 (Becquerel per cubic metre)

b. lung cancer

c. Radon sump pumps and positive-pressure ventilation systems

Chapter 8 Self-Assessment Questions

(Student textbook pages 402–3)

1. c

2. b

3. e

4. c

5. b

6. e

7. c

8. b

9. c

10. b

11. a. stainless steel **b.** seawater **c.** air

12. The excess solute will precipitate out of the solution.

13. Insoluble is not a literal term. It means less than 0.1 g of a solute will dissolve in 100 mL of solvent.

14. Water is such a good solvent because its polar oxygen-hydrogen bonds and molecular shape make it “like” many substances.

15. a. Yes, according to Table 8.3, most potassium compounds are soluble, and most nitrate compounds are soluble.

b. No, according to Table 8.3, calcium, strontium, barium, and lead form sulfate compounds with low solubility.

c. No, according to Table 8.3, only alkali ions, hydrogen, and ammonium form soluble compounds with the carbonate anion. All other carbonate compounds have low solubility.

d. No, according to Table 8.3, hydrogen, ammonium, strontium, barium, thallium, and alkali ions form soluble compounds with the hydroxide anion. All other hydroxide compounds have low solubility.

16. Hydrogen bonding occurs between the carbon and oxygen dipole in acetone and the oxygen and hydrogen dipoles in water. The polar nature of the carbon and oxygen bonds in acetone also make it a good solvent.

17. 85 g/100 g

18. 58.3% (m/m)

19. 29 mL

20. 0.4 ppm

21. 200 L

22. 943 mL; about 2.06 L of distilled water

23. $NaCl(s)$, because higher precision would be possible. The solution should be prepared by measuring a mass of $NaCl(s)$ using a balance and then dissolving the solute in water. The correct amount of water should be added to obtain the desired concentration, based on molar concentration calculations.

24. Add water to 4.0 g of $KMnO_4(s)$ until it measures 250 mL.

25. **a.** $AgNO_3(aq)$, because its molar mass is less
    b. silver chromate (Ag_2CrO_4)

26. 205 mL

27. **a.** This would decrease the concentration of the solution.
    b. A more accurate standard solution of sodium ions could be prepared using solid solute ($NaCl(s)$).

28. Answers should show an understanding of how intermolecular forces affect solubility and how solubility might affect the containment of leachate.

Chapter 9 Reactions in Aqueous Solutions

Learning Check Questions

(Student textbook page 413)

1. Any insoluble substance, such as $AgCl(s)$, $CaCO_3(s)$, or a molecular compound such as $H_2O(\ell)$ or $CO_2(g)$ are never shown as ions in a net ionic equation.

2. A chemical equation shows the complete chemical formulas of the reactants and the products; an ionic equation shows soluble ionic substances in their dissociated form.

3. Charges, as well as atoms, cannot be gained or lost during a chemical reaction.

4. There is no reaction because all compounds that contain ammonium and sodium ions are soluble in water.

5. Flame test—the colour of the flame when a substance is burned can indicate what elements are present
   Solution colour—the colour of a solution can indicate what ions are present in the solution
   Precipitate—if a precipitate forms under specific reaction conditions, it is possible to deduce the ions that were present in the solution

6. Yes, the colours of ions are produced when their electrons absorb energy (heat within a flame test and light within a solution) and are excited to a higher energy level. When the electrons fall back to lower energy levels, energy is released as specific wavelength of light.

(Student textbook page 418)

7. $c = \frac{n}{V}$; c is the concentration in moles per litre, n is amount of a given substance (in moles), and V is volume of solution (in litres)

8. You need to know the molar ratios of the reactants and products.

9. molar mass of the solute

10. The limiting reactant is the reactant that is completely consumed during a chemical reaction, limiting the amount of product that is produced

11. The cations and anions have equal but opposite charges.

12. The calculations are based on molar ratios and amounts in moles.

(Student textbook page 424)

13. If the water is hard, you will have difficulty in producing a lather using soap, or you might observe lime build-up on a heating element. If the water is soft, you will find that soap lathers easily, and you should not have notable build-up on heating elements, dishes, or shower tiles.

14. $Ca^{2+}(aq)$ and $Mg^{2+}(aq)$ are present in greater concentrations in hard water.

15. The amount of calcium carbonate and magnesium carbonate in rocks varies in different areas; these ions dissolve in rainwater as it flows through the rocks.

16. Arsenic-containing soil was deposited as silt in a river delta. Deep wells penetrate below the layer of silt and surface water is above the silt level.

17. The ions $AsO_4^{3-}(aq)$ and $PO_4^{3-}(aq)$ have the same charge and are similar in radius. Thus, they should have similar solubility when combined with a cation.

18. The signs of dental fluorosis are streaking in the teeth, and especially brown mottling. You should suggest that the patient get his/her water tested for $F^-(aq)$, if using well water. You could also suggest using a fluoride-free toothpaste and mouthwash.

(Student textbook page 434)

19. Temporary hardness is the result of dissolved $CaCO_3(aq)$ and $MgCO_3(aq)$ and can be removed by boiling water. Permanent hardness results from dissolved $CaSO_4(aq)$ and $MgSO_4(aq)$ and can only be removed by chemical means.

20. The salt is used to regenerate sodium ions on the resin beads inside the unit, after the beads have become coated with calcium or magnesium ions.

21. Oil can be supplied to both plants through the same pipeline. Also, waste heat from the electrical generating plant can supply some of the heat needed by the desalination plant.

22. Canada has sufficient supplies of fresh water.

23. There has been a steady increase in life expectancy since 1800, as water treatment techniques have improved.

24. filtering through a coarse screen to remove large particles, flocculation for removing suspended particles, filtering through graded gravel and sand, saturation with oxygen to remove volatile organic compounds, treatment with chlorine to kill harmful bacteria

Caption Questions

Figure 9.4 (Student textbook page 411): Sample answer:
red—Li^+ and/or Sr^{2+}
green—Cu^{2+} and/or Ba^{2+}
orange—Ca^{2+} and/or Na^+

Figure 9.6 (Student textbook page 412): $CuCl_2$ is soluble, while AgCl is not.

Figure 9.8 (Student textbook page 422): Over 99.98% of Earth's water is unavailable for human consumption.

Figure 9.18 (Student textbook page 431): Benefits of a home water softener—reduced deposits on heating elements, better soap lather
Disadvantages—cost, maintenance, people on low-sodium diets cannot use a water softener

Figure 9.21 (Student textbook page 433): The 1900s had the greatest increase in life expectancy.

Section 9.1 Review Questions

(Student textbook page 414)

1. A spectator ion is one that is present in aqueous solution, but does not take part in a chemical reaction. Spectator ions are often ions with small charge and large radius.

2. **a.** $Cl^-(aq)$ and $NH_4^+(aq)$ **b.** $NO_3^-(aq)$ and $Ba^{2+}(aq)$
   c. $Na^+(aq)$ and $Cl^-(aq)$

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

4. **a.** copper(II) carbonate, $CuCO_3(s)$
   b. $Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$
   c. $Na^+(aq)$ and $SO_4^{2-}(aq)$

5. **a.** Sample answers: $BaCl_2(aq)$, $BaBr_2(aq)$, $BaI_2(aq)$, or $Ba(OH)_2(aq)$; and $Na_3PO_4(aq)$, $K_3PO_4(aq)$ or $(NH_4)_3PO_4(aq)$
   Many combinations containing one substance with Ba and one with PO_4 would yield $Ba_3(PO_4)_2$.
   b. Sample answers: $MgCl_2(aq)$, $MgBr_2(aq)$, $MgI_2(aq)$, or $MgSO_4(aq)$; and NaOH(aq), or KOH(aq)
   Many combinations containing one substance with Mg and one with OH_2 would yield $Mg(OH)_2$.
   c. Sample answers: $Al(NO_3)_3(aq)$, or $Al(CH_3COO)_3(aq)$; and $Na_2Cr_2O_7(aq)$ or $K_2Cr_2O_7(aq)$
   Many combinations containing one substance with Al_2 and one with Cr_2O_7 would yield $Al_2(Cr_2O_7)_3$.

6. For each reaction shown in question 5, there are several combinations of aqueous (soluble) reactants, containing the ions shown in the question, that will yield the precipitate shown and a soluble compound as products. The soluble product and the ions that combine to produce it will not appear in the net ionic equation.

7. Answers should illustrate the steps outlined on page 408 of the student textbook.

8. Qualitative analysis determines what ions are present in a solution while quantitative analysis determines how much of a given ion is present in a solution.

9. A chemist might need to carry out qualitative analysis on a solution to test drinking water for harmful ions such as $Pb^{2+}(aq)$, $As^{3+}(aq)$, etc.

10. You could grind up the tablet and perform a flame test to confirm the presence of lithium, which would produce a crimson-red flame. Then you could perform solubility tests. You could use a precipitating solution that contains a soluble compound as a reactant, but is composed of a cation that forms an insoluble compound with carbonate. Similarly, you could use a precipitating solution containing a soluble compound, but composed of an anion that forms an insoluble compound with lithium.

11. a. $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(\ell)$
$Ca^{2+}(aq) + 2OH^-(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(\ell)$
b. Qualitative because you are looking for the appearance of the milky-white precipitate to indicate that carbon dioxide was present. The analysis consists of visual observation, not measurement of the product.

12. a. $Pb^{2+}(aq)$, $PbI_2(s)$
b. $Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$

13. From left to right:
bluish colour—chromium(II), $Cr^{2+}(aq)$, or copper(II), $Cu^{2+}(aq)$
very faint yellowish colour—iron(III), $Fe^{3+}(aq)$
yellowish-gold colour—chromate, $CrO_4{}^{2-}(aq)$
purplish colour—permanganate, $MnO_4{}^-(aq)$, or possibly cobalt(II), $Co^{2+}(aq)$ or manganese(II), $Mn^{2+}(aq)$, which both give a pink solution
greenish colour—chromium(III), $Cr^{3+}(aq)$, copper(I), $Cu^+(aq)$, iron(II), $Fe^{2+}(aq)$, or nickel, Ni^{2+}.
It would be reasonable for students to say they lack confidence in their identifications, especially for the second tube, which appears almost clear, and the last tube, which appears almost bluish-green. Precipitation reactions could be used to check inferences.

14. a. $Na_2SO_4(aq)$
b. Any solution containing a halide such as sodium chloride, NaCl(aq), sodium bromide, NaBr(aq), or sodium iodide.
c. Add $Na_2SO_4(aq)$. To the filtrate, add NaCl(aq). To the filtrate, add KOH(aq).

Section 9.2 Review Questions

(Student textbook page 421)

1. 0.10 mol/L $MgCl_2(aq)$; There are 2 mol of chloride ions for every 1 mol of magnesium chloride.
2. a. 0.5 mol/L b. 0.45 mol/L
3. 42 mL
4. 4.00×10^2 mL
5. 135 mL
6. a. 0.66 g $H_2(g)$ b. 1.5 mol/L
7. a. $H_3C_6H_5O_7(aq)$ (citric acid) b. 0.238 g
8. $PbSO_4$; 2 g
9. 2×10^{-3} mol/L
10. 2.00×10^2 mL $CaCl_2(aq)$ and 99.9 mL $K_2CO_3(aq)$
11. Measure the mass of the white powder. Add water to the powder to dissolve the alkali metal carbonate. Filter off the insoluble $MgCO_3(s)$; dry and measure the mass of the insoluble $MgCO_3(s)$. Subtract the mass of the insoluble $MgCO_3(s)$ from the mass of the original white powder to determine the mass of the alkali metal carbonate. Use the mass measurements to determine the percent by mass of alkali metal carbonate in the mixture. Use the solution to perform a flame test to identify the unknown alkali metal.
12. 0.64 mL
13. You must carefully measure the volume of solution. Then add a soluble substance that will produce an insoluble precipitate. The precipitate must be fully recovered by filtration. The filtrate must be dried before determining its mass.

Section 9.3 Review Questions

(Student textbook page 429)

1. Naturally occurring materials in the air or ground (e.g., calcium carbonate from limestone; pollutants (e.g., mercury from fluorescent tubes); chemicals added for water treatments (e.g., chlorine from municipal water treatment plant)
2. (1) Check how easily the water forms a soap lather
(2) Measure calcium carbonate content with a test strip
3. Point sources could a coal-burning power plant or the discharge of untreated sewage from a pipe. Non-point sources could include run-off from cottages into a lake or the exhaust from automobiles within the city.
4. Sediment contains minerals eroded from mountains. If the minerals include ions that endanger human health, such as arsenic and fluorine, these ions will enter the water supply.
5. Sample answer: paint, pesticide, fertilizer, pharmaceuticals

6. Answers could include a chemical reaction with a substance that will precipitate arsenic out of the water, or a distillation process. Bangladesh is a very poor country, so its government and communities cannot afford to treat water on a large scale. The effects of chronic arsenic exposure appear gradually over a long period; more immediate problems such as starvation have priority given the limited resources available.
7. No, it is a non-point source because the pollution enters the environment at many different locations.
8. The water may not be treated, and could contain a variety of pollutants. Of most concern would be micro-organisms, dissolved fertilizers (nitrates and phosphates) and pesticides.
9. dioxin and furan. Answers should include mention that these pollutants can result from burning plastics, thus people should not include plastics when burning waste.
10. The benzene level would be of most concern, since it exceeds its maximum allowable concentration, above which it is known or suspected to affect human health. The high iron level is less of a concern, since iron is judged according to its aesthetic objective. The water may look yellowish, but it is not known to have adverse health effects besides the aesthetic objective.
11. **a.** Phosphates and nitrates cause excessive growth of plants and algae.
    b. It could be the result of drainage from fertilizer used on a farm or a golf course; industrial and commercial detergents.
    c. Bacteria decomposing dead plants remove oxygen from the water. This process will occur faster in warmer weather. Also, the solubility of oxygen is lower in warmer water.
12. Answers should incorporate all of the data in the table, be clearly labelled or include a legend.
13. **a.** Hard, due to $Ca^{2+}(aq)$ and $Mg^{2+}(aq)$.
    b. The surface area of the lake is small, so wind action cause little agitation of the water. The upper layer, which contains dissolved oxygen, does not mix with the lower, cold dense water.
14. A more stable herbicide will last longer and less may be required to achieve good results. On the other hand, a stable herbicide will take longer to break down and it could enter the food chain and cause harm.

Section 9.4 Review Questions

(Student textbook page 436)

1. **a.** kettle, coffee maker, water heater
   b. Add white vinegar to the appliance, which will react with the scale to form soluble compounds that can be rinsed away.
2. Heating can trigger the precipitation of calcium carbonate from hard water.
3. Some chemicals are involved in reactions that can help remove suspended materials or kill bacteria.
4. Two sodium ions are exchanged for each calcium ion. The charges must balance, therefore $2 \times Na^{+}(aq)$ are required for each $Ca^{2+}(aq)$.
5. Oil was cheap when the plant was built, and a solar plant would require a very large surface area to collect enough energy to boil a large volume of water.
6. concentrated brine and smokestack emissions from burning fossil fuel
7. Yes, you would not have to supply salt to soften water.
8. A coarse screen removes large waste. Sand and gravel beds remove much of the suspended solids.
9. Aeration involves spraying water into the air. It helps to remove iron, volatile organic compounds, and gases from the water.
10. chlorination to kill bacteria; the addition of lime and alum; the addition of sodium fluoride.
11. The potential threat would be pollution from metal ions, and also the effect of ions such as $NH_4^{+}(aq)$ and $NO_3^{-}(aq)$ increasing plant growth and leading to a reduction of dissolved oxygen.
12. The water to be purified is added to the blackened trough. The black surface helps to absorb energy from the sun, causing water to vaporize. Water vapour condenses on the glass sheets and runs down into the troughs.
13. The salt solution prevents the growth of micro-organisms that would otherwise spoil the food. The concentration of salt draws water from the cells of the micro-organisms by osmosis, preventing the micro-organisms from multiplying, or even killing them.
14. Human waste, nitrates; detergents, phosphates
15. Any experiment should include a control. Boiling samples of water is probably the easiest approach. Note that the manufacturer does not claim to remove calcium ions, only to "alter" them.

16. **a.** Many anions will work. Sample answer: NaCl(aq); Na_2SO_4(aq)
b. Sample answer: $Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$; $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

Practice Problems

(Student textbook page 410)

1. $3Ba^{2+}(aq) + 2PO_4^{3-} \rightarrow Ba_3(PO_4)_2(s)$
2. $Sr^{2+}(aq) + SO_4^{2-}(aq) \rightarrow SrSO_4(s)$
3. $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$
4. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$; $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$
5. The precipitate is FeS(s). The spectator ions are Na^+(aq) and SO_4^{2-}(aq). The net ionic equation is $Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$.
6. **a.** Spectator ions—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. Spectator ions—Li^+(aq) and NO_3^-(aq); net ionic equation—$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(\ell)$
c. No spectator ions; $2H^+(aq) + SO_4^{2-}(aq) + Ba^{2-}(aq) + 2OH^-(aq) \rightarrow BaSO_4(s) + 2H_2O(\ell)$
7. $Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$
8. Sample answer: $Al(NO_3)_3$(aq) or $Al(CH_3COOH)_3$(aq); and $Na_2Cr_2O_7$(aq) or $(NH_4)_2Cr_2O_7$(aq); ; or Al^{3+}(aq) and $Cr_2O_7^{2-}$(aq)
9. $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$; The spectator ions are NO_3^-(aq) and K^+(aq).
10. **a.** $Pb(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow PbCO_3(s) + 2NaNO_3(aq)$
$Pb^{2+}(aq) + CO_3^{2-}(aq) \rightarrow PbCO_3(s)$
b. $Co(CH_3COO)_2(aq) + (NH_4)_2S(aq) \rightarrow CoS(s) + 2NH_4CH_3COO(aq)$
$Co^{2+}(aq) + S^{2-}(aq) \rightarrow CoS(s)$

(Student textbook page 417)

11. $NH_4^+(aq) = 0.04$ mol/L; $PO_4^{3-}(aq) = 0.01$ mol/L
12. 0.11 mol/L
13. $2Ag^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3(s)$; 0.239 mol/L
14. 1.10 g FeS(s)
15. 2.7 g
16. 24 g
17. 1.22 g
18. 0.0370 mol/L
19. **a.** 0.214 mol/L **b.** 12.5 g
20. **a.** 0.40 mol/L **b.** 16 g

(Student textbook page 420)

21. 12 g
22. 0.518 g
23. 96.4 mL
24. 0.662 g
25. 4.11 g; The calculation assumes PbI_2(s) is completely insoluble, whereas some lead(II) iodide will remain in solution.
26. 33.5 mL
27. 0.826 mol/L
28. 19.5 mL
29. 0.500 g
30. 0.0150 mol/L

Chapter 9 Review Questions

(Student textbook pages 447–9)

1. b
2. b
3. c
4. d
5. a
6. b
7. c
8. e
9. **a.** A complete chemical equation shows the complete chemical formulas of the reactants and the products; a net ionic equation shows the soluble ionic substances in their dissociated form excluding the spectator ions.
b. It shows only the ions that participate in a reaction.
10. Na^+(aq), SO_4^{2-}(aq)
11. **a.** $Cu(s) + Fe^{2+}(aq) \rightarrow Cu^{2+}(aq) + Fe(s)$
b. $Sr^{2+}(aq) + SO_4^{2-}(aq) \rightarrow SrSO_4(s)$
c. $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$
12. **a.** Cu^{2+}(aq)
b. Ca^{2+}(aq)
c. Na^+(aq) and MnO_4^-(aq)

13. Sample answer: qualitative analysis could be used to identify the presence of a chemical change (colour, gas, precipitate). Quantitative analysis would be needed to identify the products of such a change or measure the reactants/products.

14. Benefit—reduces fertilization of plant life and subsequent eutrophication of waterways
Drawback—cost, or difficulty of disposing of solid waste

15. From the atmosphere as a result of burning fossil fuels; run-off from fields recently fertilized or sprayed with pesticide; incomplete waste-water treatment

16. Aquatic life is at risk because $NO_3^-(aq)$ increases plant growth, leading to a reduction of dissolved oxygen in waterways. Nitrate ion is also responsible for "blue baby" syndrome.

17. 0.010 mol

18. 1.5 mol/L Mg^{2+}; 3.0 mol/L NO_3^-; One formula unit of magnesium nitrate dissociates into one magnesium ion and two nitrate ions.

19. a. Magnesium hydroxide is not very soluble, so the compound must be taken as a suspension.

b. $Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(\ell)$;
$Mg(OH)_2(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(\ell)$

20. Molar masses—$Na_2CO_3 = 105.99$ g/mol;
$MgCO_3 = 84.32$ g/mol

Each type of solution gives 1 mol of $CO_3^{2-}(aq)$ per mole of dissolved compound. The dissolved masses are the same, therefore the greater amount (in moles) of dissolved compound is that with the smallest molar mass. The solution of magnesium carbonate has the greater concentration of carbonate ions.

21. 0.015 mol/L

22. a. $Na^+(aq) = 5.91$ mol/L; $NO_3^-(aq) = 5.91$ mol/L

b. $Ca^{2+}(aq) = 0.88$ mol/L; $CH_3COO^-(aq) = 1.8$ mol/L

c. $NH_4^+(aq) = 1$ mol/L; $PO_4^{3-}(aq) = 0.4$ mol/L

23. 0.22 g $CaCO_3(s)$

24. $Zn(s) + Pb^{2+}(aq) \rightarrow Pb(s) + Zn^{2+}(aq)$; 0.096 mol/L

25. $Cu^{2+}(aq) + 2F^-(aq) \rightarrow CuF_2(s)$; 1.46 mol/L

26. a. $MgCl_2(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + 2NaCl(aq)$

b. $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$

c. 0.9 g

27. Sample answer: Sodium, strontium, and barium. Colour is largely subjective, so without definitive colour samples rather than descriptions, I cannot be certain. For example, is the middle colour crimson or bright red, or does the little line of orange in the flame make this reddish-orange?

28. (1) Flame test; calcium would cause a red-orange flame and potassium would cause a lavender flame
(2) add sodium carbonate solution to each test tube; calcium chloride will give a white precipitate of calcium carbonate and potassium nitrate will give no precipitate

29. Reducing water used—turn the tap off while brushing teeth; installing low volume toilets, or displacing water in a conventional toilet tank with a brick; low flow showerhead; watering gardens and lawns only as necessary
Reducing or eliminating pollution—do not flush household wastes or pharmaceuticals down the toilet; do not allow kitchen grease down the sink; do not pour engine oil, pesticides, etc. into storm drains

30. Answers should include the use of students' own words to demonstrate a thorough understanding of the functioning of the components of reverse-osmosis equipment. There are good illustrations of the functioning of this equipment available on the Internet. Diagrams should be clearly labelled.

31. Calculate the mass of a sample of pure $BiCl_3(s)$ and calculate the amount (in moles). Add excess water to make sure all the $BiCl_3(s)$ reacts. Filter and dry the solid. Mass the dry precipitate and calculate the amount (in mol) based on each of the three possible products. Compare with the amount expected based on the chemical equation and the amount of $BiCl_3(s)$ used.

32. Answers should be based on research from several different appropriate and credible sources. The table making comparison should be clearly labelled.

33. Answers should be based on research from several different appropriate and credible sources. Use of students' own words in the letters should demonstrate a thorough understanding of the processes. Recommendations should be supported with information from research.

34. Assessment Checklist A-13: Concept Map can assist students in creating an organizer that includes points summarized on page 446.

35. 0.3 mol/L

36. a. run-off from cattle waste into a nearby stream; untreated sewage

b. $(NH_2)_2CO\ (aq) + H_2O(\ell) \rightarrow 2NH_3(aq) + CO_2(g)$

$NH_3(aq) + 3O_2(g) \xrightarrow{\text{oxidizing bacteria}} 2NO_2^-(aq) + 2H^+(aq) + 2H_2O(\ell)$

$2NO_2^-(aq) + O_2(g) \xrightarrow{\text{oxidizing bacteria}} 2NO_3^-(aq)$

37. Wild game could absorb some lead from shotgun shells. This might occur directly when game is shot, and/or from spent pellets being present in water systems; the lead builds up in the animal's body over time. Lead passes into the body of whoever eats the animal. Evidence could be gathered by comparing lead levels in people who consume wild game with those who do not, or who eat less.

38. Untreated water will contain pollution from roads, parking lots, fields, etc. Run-off will contain nitrates, phosphates, pesticides, oil, and bacteria. A common result is that people are warned not to swim in a lake following a rain storm.

39. a. 4 L

b. To ensure that all of the chlorine present in the sample actually reacts

Chapter 9 Self-Assessment Questions

(Student textbook pages 450–1)

1. c

2. e

3. a

4. d

5. b

6. d

7. a

8. b

9. e

10. d

11. (1) Add a solution containing $Cl^-(aq)$. A white precipitate indicates the presence of $Ag^+(aq)$; no precipitate forms in the solution containing $Ca^{2+}(aq)$. (2) Flame tests—Ag gives no colour; Ca gives a reddish-orange flame

12. a. $2Na_3PO_4(aq) + 3Ca(OH)_2(aq) \rightarrow 6NaOH(aq) + Ca_3(PO_4)_2(s)$

b. $3Ca^{2+}(aq) + 2PO_4^{3-}\ (aq) \rightarrow Ca_3(PO_4)_2(s)$

13. a. tin(II) phosphate, $Sn_3(PO_4)_2(s)$

b. $K^+(aq)$ and $Cl^-\ (aq)$

c. $3Sn^{2+}(aq) + 2PO_4^{3-}\ (aq) \rightarrow Sn_3(PO_4)_2(s)$

14. $Cr^{2+}(aq)$, $Cu^{2+}(aq)$; A bluish flame in a flame test would show that the solution contains copper(II) instead of chromium(II). Also, hydrogen sulfide could be added: copper(II) will form a precipitate with the sulfide ions.

15. Sample answers: Use NaCl(aq): $Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$; Use $Na_2SO_4(aq)$: $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$

16. 5.6 g

17. a. 2.0 mol **b.** 3×10^3 kg

18. 25 g

19. Ground water lies beneath Earth's surface. Surface water is collected on Earth's surface. depends on location

20. lakes, rivers and reservoirs.

21. Temporary hardness can be removed by boiling water. Permanent hardness cannot be removed by boiling water. Temporary hardness is caused by dissolved $CaCO_3$ and $MgCO_3$. Permanent hardness is caused by dissolved $CaSO_4$ and $MgSO_4$.

22. Soft water does not contain the $Ca^{2+}(aq)$ and $Mg^{2+}(aq)$ that combine with soap molecules, thus making it more difficult to rinse the soap away. Also, it is easy to use more soap than required to form a lather using soft water. In this case, there is excess soap to rinse.

23. The water was not treated and contained micro-organism.

24. The common methods are distillation and reverse osmosis. Each is energy intensive. Distillation is achieved by heating saline water using solar energy or a non-renewable fuel such as oil or gas. The solid salt must be disposed and may contribute to further environmental problems. Reverse osmosis uses energy to force water through a semi-permeable membrane. This results in salt-free water on one side of the membrane, and a more concentrated saline solution on the other side, which must be disposed of. The advantage of each method is the production of potable water in areas where the supply is greatly limited. The advantage of reverse osmosis over distillation is that reverse osmosis is more energy efficient and its does not produce any thermal pollution.

25. Microorganisms such as bacteria and viruses; dissolved metals and radioactive elements from ground rock; fluoride ion. Bacteria and nitrates from human waste; phosphates from soaps and detergents; fertilizers and pesticides from lawns, gardens and agricultural fields; industrial waste; leachates from landfills and waste dumps.

Chapter 10 Acids and Bases

Learning Check Questions

(Student textbook page 457)

1. An Arrhenius acid is a substance that ionizes in water to produce one or more hydrogen ions, $H^+(aq)$. An Arrhenius base is a substance that dissociates in water to form one or more hydroxide ions, $OH^-(aq)$.
2. **a.** Sample answer: citrus fruit, tomatoes, vinegar, carbonated drink
   b. Sample answer: soap, detergent, ammonia solution (window cleaner), oven cleaner
3. Answers should include similarities, such as conduct electricity and corrode tissues. Properties exclusive to acids should include sour taste, pH less than 7, turn litmus red, produce no colour change in phenolphthalein, corrode metals, react with active metals to produce hydrogen gas, and react with carbonates to produce carbon dioxide gas. Properties exclusive to bases should include bitter taste, slippery texture, pH greater than 7, turn litmus blue, turn phenolphthalein pink, do not corrode metals, do not react with active metals, and do not react with carbonates.
4. **a.** acid **b.** base **c.** acid **d.** acid
   e. base **f.** acid **g.** base **h.** base
5. The paper will remain blue, because the solution is basic.
6. **a.** The $H^+(aq)$ concentration is higher than the $OH^-(aq)$ concentration.
   b. The $H^+(aq)$ concentration is higher than the $OH^-(aq)$ concentration.
   c. The H^+ (aq) concentration is lower than the OH^- (aq) concentration.

(Student textbook page 462)

7. A strong acid is one that ionizes completely into ions; for example, $HCl(aq)$. A concentrated acid is one with a relatively large amount of acid dissolved in the solution; for example, 12 mol/L $HCl(aq)$.
8. **a.** 0.01 mol/L $NaOH(aq)$ **b.** 4 mol/L $HF(aq)$
9. Both acids are weak. Reasoning: Soft drinks are consumed by humans and it is dangerous and deadly to consume strong acids. Also, neither is on the list of strong acids.
10. The acids are listed in order of increasing strength. The addition of more oxygen atoms increases the polarity of the bond between the ionizable hydrogen atom and the oxygen atom it is attached to.
11. Diagrams should be similar to Figure 10.9 on page 461. They should show a high degree of dissociation in a strong base, compared with a high concentration of solute in a concentrated base, and a low degree of dissociation in a weak base, compared with a low concentration of solute in a dilute base.
12. The safety hazards associated with strong acids and bases are far greater than those associated with weak acids and bases. Strong acids and bases are highly corrosive and should never be consumed or allowed to come in contact with skin, while weak acids and bases are actually ingredients in some common foods and beverages

(Student textbook page 467)

13. A salt and water are produced by a neutralization reaction.
14. The pH of the equivalence point in a titration between a strong acid and a strong base is 7.0, while the equivalence point for titrations involving either a weak base or a weak acid is slightly below 7.0 or slightly above 7.0 respectively.
15. If both concentrations were unknown, you would have no way to calculate the concentration of either the base or the acid. The volume and concentration of the known solution and the volume of the unknown solution at the equivalence point are all needed to calculate the unknown concentration.
16. The term *neutralization reaction* refers to a reaction between an acid and a base. It does not mean that all of the hydrogen ions and all of the hydroxide ions have been neutralized.
17. The end point occurs when the indicator changes colour. The equivalence point occurs when stoichiometric quantities of acid and base have been mixed together.
18. Phenolphthalein changes colour in basic solution between a pH of 8.2 and 10.0. Thus, more of the weak base will be added than required for equivalence.

Caption Questions

Figure 10.1 (Student textbook page 454): Both lemons and grapes are acidic.

Section 10.1 Review Questions

(Student textbook page 463)

1. a. In the Arrhenius theory, a substance that does not contain hydrogen atoms cannot be an acid because it could not ionize to form $H^+(aq)$.

b. A substance that does contain hydrogen atoms may not be an acid if the substance does not ionize in water to form $H^+(aq)$.

2. The properties of an acid are due to the presence of $H^+(aq)$. $H_2SO_4(aq)$ ionizes to form $H^+(aq)$. Glucose is a covalent compound that does not ionize in water.

3. Ionization diagrams should show a covalently-bonded compound breaking apart in water to form ions. Dissociation diagrams should show an ionic compound breaking into separate ions when dissolved in water.

4. A universal indicator is a solution (or paper treated with the solution) that changes to specific colours throughout the pH range.

5. Test with an indicator (e.g., blue litmus turns red), and mix the unknown compound with a carbonate to see if $CO_2(g)$ forms.

6. A is basic, B is neutral, C is acidic

7. Marble is a form of calcium carbonate, and vinegar is acidic. Acids react with carbonates, thus vinegar will corrode marble.

8. a. $pH < 7$ **b.** $pH > 7$

9. a. coconut milk, pH 6.5; buttermilk, pH 4.5; orange juice, pH 4.0; cranberry juice, pH 2.4; lime juice, pH 2.2

b. lime juice, cranberry juice, orange juice, buttermilk, coconut milk; The lime juice will taste the most sour because it is the most acidic. The other beverages will taste progressively less sour according to their acidity.

10. $H_2O(\ell)$, $H^+(aq)$, and $Cl^-(aq)$ are present in an aqueous solution of hydrochloric acid. Hydrochloric acid is a strong acid and thus completely ionizes in water, leaving only H^+ and Cl^- ions. Concentrated hydrochloric acid is normally 12.0 mol/L, so a 1 mol/L solution must have been diluted with water.

11. a. Strong Arrhenius acids: $H_2SO_4(aq)$ and $HCl(aq)$

b. Aqueous ammonia (NH_3) is a weak Arrhenius base.

12. It means very few of the particles making up the acid ionize to form $H^+(aq)$.

13. $NH_3(aq)$, $CH_3COOH(aq)$, $HCl(aq)$; Ammonia is a weak base, acetic acid is a weak acid, and hydrochloric acid is a strong acid.

14. A conductivity tester can be used to distinguish between a strong base and a weak base. The conductivity tester will glow brightly when testing the strong base, but will glow only faintly when testing the weak base.

15. Strong acids are corrosive and can harm human tissue. Weak acids, on the other hand, are not as corrosive, and people can handle these acids and even consume them without being harmed.

16. Yes, it is possible. The term *concentrated* refers to the amount of solute in solution, and the term *weak* means that the acid only partly ionizes in solution. So a concentrated solution of a weak acid will have a large amount of acid dissolved, but very few ions will form in it.

Section 10.2 Review Questions

(Student textbook page 470)

1. acid + base → salt + water

2. a. $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(\ell)$

b. $2HI(aq) + Mg(OH)_2(aq) \rightarrow MgI_2(aq) + 2H_2O(\ell)$

3. At equivalence, equal amounts of acid and base have reacted together. The end point of an indicator refers to the pH range over which the indicator changes colour. Sometimes the end point for an indicator is outside of the pH range where equivalence happens.

4. The mistake results in the acid being diluted in the pipette. Therefore, less base will be needed to reach the end point and the concentration of the acid will be calculated to be smaller than it really is.

5. The mistake causes extra base to be added, therefore requiring more acid from the burette. The calculated base concentration will be greater than it really is.

6. Shampoo is basic. Rinsing with vinegar neutralizes any shampoo left in the hair. In addition, vinegar is a weak acid. Thus, too much vinegar does not leave the hair overly acidic.

7. At equivalence, a fraction of a drop of base will make the solution basic enough to change the colour of the indicator.

8. 0.5 mol

9. Milk of magnesia contains a base. The fluids in the stomach contain an acid. The base in the milk of magnesia reacts with this acid (a neutralization reaction), thereby reducing the irritation.

10. 21.3 mL

11. The metal cation from the base combines with the anion from the acid to form a salt.

12. Sample answer: (1) Rinse a pipette with distilled water and then twice rinse the same pipette with strontium hydroxide.
(2) Using this pipette, transfer a specified volume (e.g., 25 mL) of strontium hydroxide into an Erlenmeyer flask.
(3) Add two or three drops of phenolphthalein to the flask.
(4) Rinse a burette with distilled water and then twice rinse the same burette with 0.250 mol/L sulfuric acid.
(5) Clamp the burette to a retort stand and add the sulfuric acid to the burette.
(6) Add the sulfuric acid from the burette to the flask and stop when the solution in the flask has a permanent colour change.
(7) Use the volume and concentration of sulfuric acid to calculate the concentration of strontium hydroxide (see the Sample Problem on page 465).
The equation for the reaction is
$H_2SO_4\,(aq) + Sr(OH)_2\,(aq) \rightarrow SrSO_4\,(s) + 2H_2O\,(\ell)$

13. 6.554×10^{-2} mol/L

14. The volume of acid will be the same. The net ionic equation for each reaction is identical, and the concentration of the acids is the same. Thus, the same volume of acid will be required to reach equivalence.

15. In everyday language, *salt* refers only to sodium chloride, NaCl(s). In chemistry, *salt* refers to any ionic compound formed during a neutralization reaction by combining the cation from a base and the anion from an acid.

16. The HCl should go in the burette and the NaOH will be measured with the pipette. The titrant, which is the solution that goes in the burette, is the solution with the known concentration. A set volume of the unknown solution is measured out with a pipette.

Practice Problems

(Student textbook page 466)

1. 2.12 mol/L
2. 67.5 mL
3. 0.1298 mol/L
4. 107 mL
5. 0.32 mol/L
6. 0.128 mol/L
7. 2×10^{-4} mol/L
8. 87.3 mL
9. 118 mL
10. 2.5×10^{1} mL

Chapter 10 Review Questions

(Student textbook pages 479–81)

1. d
2. a
3. c
4. c
5. a
6. b
7. d
8. e
9. Sample answer: sodium hydroxide, NaOH(aq), potassium hydroxide, KOH(aq), ammonia, NH_3(aq), magnesium hydroxide, $Mg(OH)_2$(aq)
10. Sample answers:
    a. hydrochloric acid, HCl(aq), hydroiodic acid, HI(aq)
    b. sulfuric acid, H_2SO_4(aq), carbonic acid, H_2CO_3(aq)
    c. phosphoric acid, H_3PO_4(aq), citric acid, $H_3C_6H_5O_7$(aq)
11. The acidic properties of acids, such as HCl(aq), and the basic properties of bases containing OH, such as NaOH, as listed in Table 10.1 are observations that can be explained using the Arrhenius theory.
12. a. $Ca(OH)_2(aq) + 2HCl(aq) \rightarrow 2H_2O(\ell) + CaCl_2(s)$
    b. $2H_3PO_4(aq) + 3Sr(OH)_2(aq) \rightarrow 6H_2O(\ell) + Sr_3(PO_4)_2(s)$
    c. $2NaOH(aq) + H_2SO_4(aq) \rightarrow 2H_2O(\ell) + Na_2SO_4(s)$
13. Acids taste sour and bases taste bitter. Slightly sour or even very sour food tastes good to people but bitter foods generally are not enjoyed.
14. Perchloric acid has two more oxygen atoms than chlorous acid does. The more oxygen atoms in a molecule, the greater the polarity of the bond between the hydrogen atom and the oxygen atom that it is attached to. So the hydrogen atom of perchloric acid is more likely to ionize than the hydrogen atom of chlorous acid.

15. Answers should include:
Definition—a substance that dissociates in water to form one or more hydroxide ions, $OH^-(aq)$
Characteristics—based on early acid-base theory developed by Swedish chemist and physicist Arrhenius; must contain a hydroxyl group, which is a source of hydroxide ions in an aqueous solution; undergoes dissociation in water; mostly ionic compounds.
Examples—$NaOH(aq)$, $KOH(aq)$
Non-examples—$HCl(aq)$, $HBr(aq)$

16. Quantitative analysis the name given to any process that determines how much of a substance is present. Titration is used to determine the concentration of an acid or a base by reacting it with a certain amount of a base or acid that has a known concentration.

17. You could use a pH meter to monitor the change in pH. The equivalence point is the point at which pH is 7. So you would add base to the acid until the pH meter reads 7.

18. $HOCl(aq) \rightarrow H^+(aq) + OCl^-(aq)$

19. Test each solution with litmus paper. The acidic solution will turn blue litmus red; red litmus will stay red. The basic solution will turn red litmus blue; blue litmus will stay blue. The neutral solution will not change the colour of either type of litmus paper.

20. 0.101 mol/L

21. a. pink
b. colourless
c. colourless
d. pink

22. pH 4.4–6.0; Phenolphthalein will be colourless.

23. 0.5 mol

24. a. 10.90 mL, 10.05 mL, 9.95 mL
b. Answers should use only the data from the last two trials. The first trial used much more $NaOH(aq)$ than the last two trials indicating that the student probably over-titrated that sample.
c. 0.21 mol/L

25. Answers must include determining the pH of the solutions (using either indicators or a pH meter), and comparing the conductivity of the solutions.

Solution	pH	Conductivity
strong base	largest	high
strong acid	lowest	high
weak base	>7	low
weak acid	<7	low
neutral ionic	7	high
neutral molecular	7	very low

26. e, a, d, c, f, g, b

27. a. An acid-base reaction results in the *neutralization* of the acidic properties of the acid and the basic properties of the base. In a neutralization reaction, only some of the acidic and basic properties may be neutralized. However, in everyday language, *neutralization* means that something is completely neutralized.
b. The solution will only be neutral (pH 7) if both acid and base were strong.

28. Wet beakers will dilute the solutions put into them. A wet Erlenmeyer will also dilute a solution placed into it, but in this case the solution contains a specific amount of reactant.

29. Hydrochloric acid is a strong acid, which means that it all the molecules will ionize in solution. Acetic acid is a weak acid, which means that very little of the molecules will ionize in solution. So a 1.0 mol/L solution of hydrochloric acid contains 1.0 mol/L hydrogen ions. A 1.0 mol/L solution of acetic acid contains much less than 1.0 mol/L hydrogen ions. Because the hydrochloric acid has so many more hydrogen ions in solution, it will react more violently with the magnesium than the acetic acid will.

30. The first diagram should show H-Br and K-O-H bonded together. The second diagram should show H^+, Br^-, K^+, and OH^- ions. The final diagram should show H-O-H (water), and K-Cl bonded together.

31. The diagram shows a model of a dilute solution of a strong acid. It is a strong acid because all of the molecules are ionized. It is dilute because there are not very many ions in solution.

32. You cannot determine which is more dangerous because the terms *concentrated* and *strong* refer to two different properties of an acidic solution. The concentrated acid could be a weak acid that is not dangerous. The strong acid could be a solution that is so dilute that it is not harmful.

33. Answers should include a comparison of the scientific and everyday definitions of a word used in this chapter and explain how the difference could cause confusion. Some words students may describe include *salt, concentrated, dilute, strong, weak,* and *neutralization.*

34. Organizers should reflect the content of the chapter as summarized on page 478. Concepts should be organized from the most general to the most specific and include the key terms of the chapter. The organizer should show multiple levels (general to specific) and valid cross links among concepts, using appropriate linking words or symbols. The organizer should have an effective title, be easy to follow, and can show prior knowledge as well as new knowledge.

35. Between 2.8 and 4.8.

36. Soap contains a base and bases have bitter tastes.

37. **a.** 8×10^1 kg (rounded to appropriate number of significant figures)
    b. 3.68 mol/L **c.** yes

38. It is better to have the acid in the Erlenmeyer flask, because the colour change will be from colourless to pink. This is easier to see, compared with a change from pink to colourless.

39. 0.429 g

40. 24.0 g/mol

41. 67.6% (m/m)

42. Presentations should be based on research from credible sources and in students' own words. Students should be able to answer questions from the audience concerning concepts.

Chapter 10 Self-Assessment Questions

(Student textbook pages 482–3)

1. b
2. a
3. b
4. e
5. e
6. b
7. c
8. b
9. b
10. d
11. A common household acid (such as vinegar or lemon juice) and a common base (such as glass cleaner, household ammonia or drain cleaner).
12. **a.** Acidic—sour taste; tend to be caustic; react with magnesium to generate hydrogen gas; turn blue litmus red; aqueous solution conducts electricity; pH of aqueous solution <7
    b. Basic—bitter taste; may be caustic; do not react with Mg; slippery to the touch; turn red litmus blue; aqueous solution conducts electricity; pH of aqueous solution >7
    c. Neutral—not caustic; does not react with Mg; no effect on litmus indicator; aqueous solution may or may not conduct electricity; pH of aqueous solution $= 7$
13. Hydrangea have a natural indicator in their flowers. When the soil that the flowers are growing in is acidic, the flowers are blue. Orange peels have citric acid in them, so adding the peels to the soil will make the soil more acidic.
14. $H_2S(aq)$ and $H_3AsO_4(\ell)$
15. $KOH(aq)$ and $Ca(OH)_2(aq)$
16. The residents should collect rain samples from different areas including areas both close to and far away from the factory. They should then test the pH of the rain samples with a pH meter or a universal indicator. Finally, they should look for trends in pH and distance from the factory.
17. Some of the known solution in the burette displaced the air bubbles in the tip of the burette. This means that some of the volume the student thought went into the Erlenmeyer actually stayed in the burette. Less of the titrant went into the flask. That means that less titrant was needed to neutralize the solution and the calculated concentration of the unknown is greater than the actual concentration of the unknown.
18. A strong acid ionizes completely into cations and hydroxide ions. Examples include any Group 1 hydroxide or $Ca(OH)_2$ does not ionize completely, so many of the original weak acid molecules remain intact in the solution. See examples on pages 458–9 in the student textbook.
19. Flowcharts should include: rinsing the pipette; filling the pipette; transferring the solution; adding the indicator; rinsing the burette; filling the burette; and titrating the unknown solution.
20. The simplest procedure would be to use a portable pH meter. Universal indicator could be used, but it will not give as accurate results. Titrations could be accurate, but would involve standardizing solutions, etc. and would be much more time-consuming. The pH will likely change most as a result of spring run-off adding acidic deposition to the stream, lowering the pH.
21. (A) strong acid, strong base, ionic salt solution
    (B) weak base, weak acid

22. 1.505

23. The pH is less than 7.

24. 0.2479 mol/L

25. 0.585 mol/L NaOH(aq)

Unit 4 Review Questions

(Student textbook pages 487–91)

1. e

2. a

3. c

4. d

5. a

6. a

7. c

8. c

9. a. A gram of salt dissolved in a litre of solution; a few mL of acetic acid in a large volume of water

b. 50 g of sugar dissolved in 100 mL water; an equal (v/v) mix of water and ethyl alcohol, or acetic acid

c. Oxygen bubbled through water; any insoluble substance (such as calcium carbonate) that is present as a solid in a solution of the substance

10. a. No effect. The solution is already saturated.

b. Most probably will increase the solubility because most solids are more soluble at greater temperature.

c. No effect. The two solutes do not react and each dissolves independently of the other.

d. No effect. Pressure has negligible effect on the solubility of a solid in water.

11. a. Probably ppm, because the concentration should be small, though it could be % (m/m) if the concentration is relatively large

b. Probably mol/L, because the quantities reacting are important

c. Probably % (v/v), because the relative volume of methane in air could be explosive; could be ppm if smaller concentrations are being monitored.

d. Probably % (m/m), because tin is a relatively large constituent of the alloy.

e. % (v/v) because the two liquids are miscible

f. ppb, because the concentration is very small

12. K^+(aq) and Cl^- (aq)

13. Nitrates, phosphates and insecticides; run-off from fields; dissolved carbon dioxide and sulfur dioxide; acidic gases resulting from burning fossil fuels. Organic liquids from business sources, such as dry cleaning solvents.

14. Test with an indicator such as litmus or phenolphthalein. In acidic solutions, litmus is red. In basic solutions, litmus is blue. Second, add a strip of magnesium or a piece of zinc to the solution. In acidic solutions, the metal will generate hydrogen gas. In basic solution, there is no reaction.

15. a. strong acid

b. weak acid

c. weak base

d. strong base

e. strong acid

f. weak base

16. The product of the reaction is a salt. Depending on the ions in the salt, the solution may be acidic, neutral, or basic. The indicator must correspond closely with the pH of the solution. For example: $HCl(aq) + NaOH \rightarrow NaCl(aq) + H_2O(\ell)$. The solution is neutral; use phenolphthalein or bromothymol blue.

$HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$. The solution is acidic; use methyl orange.

$CH_3COOH(aq) + NaOH \rightarrow CH_3COONa + H_2O(\ell)$. The solution is basic; use phenolphthalein.

17. a. The concentration will be a bit low since some sodium ions will precipitate during the preparation.

b. NaCl(s) could be used.

18. The low pressure allows gases to come out of solution in the synovial fluid, increasing the pressure on the joints.

19. You could add sulfuric acid, which would precipitate Ca^{2+} but not Mg^{2+}. $Ca(NO_3)_2(aq) + H_2SO_4(aq) \rightarrow CaSO_4(s) + 2HNO_3(aq)$

20. 3.0 mol/L

21. Assuming the density of solution is 1.0 g/mL, the atomic mass of *X* is 19 g/mol and the anion is F^-.

22. Precipitate: iron(III) hydroxide, $Fe(OH)_3(s)$; Net ionic equation: $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

23. 0.137 g

24. a. 0.3 mol/L **b.** 0.7 mol/L

25. The ferric ion, Fe^{3+} is smaller in radius and has greater charge compared with the ferrous ion, Fe^{2+}. Thus, Fe^{3+} forms stronger ionic bonds than Fe^{2+}.

26. 0.716 mol/L

27. a. molar mass KSCN = 97.19 g/mol. A small volume is required. If 25 mL of 0.1 mol/L KSCN(aq) is made, dissolve 0.24 g KSCN in 25 mL distilled water. If 50 mL of 0.1 mol/L KSCN(aq) is made, dissolve 0.49 g KSCN in 50 mL distilled water.

b. Make up solutions containing a series of known concentrations of Fe^{3+}(aq) using a soluble salt, such as $FeCl_3$(s). Add the same number of drops of KSCN(aq) to equal volumes of each reference solution and the tap water sample. Compare the colour of the tap water $Fe(SCN)^{2+}$(aq) with the reference solutions and identify the closest match.

28. The liquids mentioned will kill the bacteria that break down organic compounds in wastewater. Furthermore, many of these products are not biodegradable and will remain in the ground water thereby contaminating it.

29. The pH is less than 3.2.

30. a. 3.67 g

b. When decanting the HNO_3(aq) within which the $SrSO_4$(s) is a precipitate, some of the $SrSO_4$(s) may not be retrieved, and $SrSO_4$(s) may not be dried sufficiently before measuring its mass.

31. a. 4 L

b. To ensure that all of the chlorine reacts, and that all of the iodine formed dissolves

32. 6×10^2 mL

33. 0.886 mol/L

34. Answers should include a step to determine whether the substance is homogeneous or heterogeneous. Solutions and pure substances are homogeneous. They should also include a step to determine whether the substance changes state over a temperature range. Solutions change state over a temperature range; pure substances change state at a fixed temperature.

35. The CH_3 groupings are non-polar, and these help acetone dissolve molecular compounds. Hydrogen bonding between water molecules and the polar C=O group explains the solubility of acetone in water.

36. a. "salt free" means less than 5 mg of sodium or salt per serving

b. "low in salt" means 140 mg or less of salt per 100 g, if the food is a prepackaged meal

c. "lower in salt" means the food contains at least 25% less salt compared with a similar prepackaged food

37. Safety equipment includes gloves, goggles or face shield, an apron, and a fume hood. Add concentrated acid to water. Only two significant digits are needed, thus graduated cylinders can be used for volume measurements. Measure 41.7 mL concentrated acid and pour into about 1.5 L distilled water. Add distilled water to make up the volume to 2.0 L in a 2-L graduated cylinder.

38. Chlorine in a swimming pool kills bacteria, but too much irritates the eyes. Fluoride ion in drinking water protects against cavities, but too much causes dental fluorosis.

39. Measure the mass of a sample of the mixture. Add the mixture to boiling water to dissolve the $PbCl_2$. Use the mass of mixture to determine the volume of water. For example, if the mixture has mass 5 g, not less than 150 mL of water should be boiled. Decant the solution, and dry the solid AgCl. Measure the mass of dry AgCl, and determine the percent present: (mass AgCl/mass mixture) $\times$ 100 %

40. Very young children often swallow toothpaste. Increased fluoride consumption can lead to problems with bones and mottling of teeth.

41. a. Diagrams should illustrate each step in the water cycle and should be clearly labelled. Students should be able to explain why they selected each sources and why these sources were appropriate and credible.

b. (1) Changes of state involving evaporation and condensation.
(2) Filtration through sand and gravel.
(3) Bacterial action to decompose organic compounds.
Water treatment facilities use each of these steps, except evaporation and condensation.

42. The experiment should describe how to measure and adjust pH by adding compounds to the soil in which plants will be grown. The pH can be measured with universal pH paper, or a portable pH meter. The soil pH can be adjusted by adding inorganic compounds to it. Adding lime, $CaCO_3$(s), will increase pH. Lowering the pH could be achieved by adding acetic acid. Adding nitrogen fertilizers would also lower the pH, but the added nutrient would be an unfair test of changing only soil pH. The experiment should also describe how plant growth will be measured to compare the effect of changing pH.

43. Volume/mass percent—used to measure intravenous solutions; easy to determine from simple mass and volume measurements; not useful for calculations that involve reactions in solutions

Mass percent—used to measure concentration of metals in an alloy; easy to determine from mass and volume measurements; not useful for calculations that involve reactions in solutions

Volume percent—used to measure solutions prepared by mixing liquids (e.g., pharmaceutical solutions); easier to measure volume than mass for liquids; easy to determine from mass and volume measurements; not useful for calculations that involve reactions in solutions

Parts per million and parts per billion—used to measure safety limits for contaminants (e.g., mercury in water); convenient for very small concentrations; not useful for calculations that involve reactions in solutions

Molar concentration—used to measure solutions used as reactants; directly related to the number of solute particles in a solution; convenient for stoichiometric calculations that involve reactions in solutions; not as simple (amount in moles is needed)

44. A saturated solution of $Ca(OH)_2(aq)$ will result in a relatively small concentration of $OH^-(aq)$. Thus, using the Arrhenius definition of a base, it would be classified as weak. However, what little dissolves dissociates completely, and on this basis $Ca(OH)_2(aq)$ is usually classified as a strong base.

45. a. The graph should show an increase in solubility increasing temperature.

b. The graph should show no change in solubility with increasing pressure (i.e., the graph should be a horizontal line.)

c. The graph should show a decrease in solubility with increasing temperature.

d. The graph should show an increase in solubility with increasing pressure.

46. Posters should summarize the key points in this chapter that deal with working with acids. It should be evident that students considered the most effective way to communicate safety information to their audience.

47. Charts should outline:

Definition—a homogeneous mixture of two or more substances

Characteristics—uniform composition throughout; substances remain evenly distributed once mixed; homogeneous at microscopic scale; all samples of a given solution contain the same substances in the same proportions

Examples—cup of tea, stainless steel spoon; natural gas; antifreeze

Non-examples—beach sand, tomato sauce

48. a. Oil and water are immiscible. The oil coats marine life and birds, affecting the properties of their skin and feathers. Oil is poisonous when ingested.

b. Oil dispersants are chemical agents such as surfactants, solvents, and other compounds. They are used to reduce the effect of oil spills by changing the chemical and physical properties of the oil. By enhancing the amount of oil that physically mixes into the water, dispersants can reduce the potential that a surface slick will contaminate shoreline habitats. Their effect on marine life is a subject of continuing research, and there is a chance these dispersants are toxic.

49. 210 mL

50. 0.05 ppm

51. Molar mass $C_6H_{12}O_6 = 180.16$ g/mol

$0.0556 \text{ mol/L} \times 180.16 \text{ g/mol} = 10.02 \text{ g/L}$

100 mL requires 1.00 g

Dissolve 1.00 g in about 70 mL distilled water in a 100 mL volumetric flask. Make the volume to the calibrated line.

52. a. Pike require a greater concentration of dissolved oxygen. More oxygen dissolves in water at a cooler temperature.

b. Trout would be in deeper waters compared with pike.

53. HgS has very low solubility. Virtually all the mercury ion is removed, and once disposed of, there is very little chance of any appreciable amount dissolving into the groundwater.

54. a. 3×10^7 L

b. Steps can be outlined in narrative or clearly labelled diagram or flow chart. Students should be able to explain research choices and why sources were appropriate and credible.

c. Answers should be based on the research students conducted for b. Students should justify their answers using their own words to demonstrate a thorough understanding.

55. Shampoos are basic. Thus, after washing, the hair is basic and more brittle. Rinsing with vinegar neutralizes the shampoo and leaves the hair acidic and therefore stronger.

56. a. 0.766 mol/L

b. Molar mass HOCl = 52.46 g/mol

1 L of solution contains $52.46 \text{ g/mol} \times 0.766 \text{ mol} = 40.18$ g

percent (m/v) $HOCl(aq) = (40.18/1000) \times 100 = 4.02\%$

57. **a.** The sodium chloride pellets are used to remove calcium and magnesium ions from the water.
b. Sodium hydroxide is corrosive.

Unit 4 Self-Assessment Questions

(Student textbook pages 492–3)

1. b
2. b
3. d
4. e
5. d
6. c
7. c
8. b
9. c
10. c
11. **a.** soluble; all potassium salts are soluble
b. insoluble; most sulfides are insoluble
c. insoluble; the solubility guideline table lists Cl^- ion as insoluble with Pb^{2+}
d. insoluble; large charge on each ion
12. $I_2(s)$ is a covalent molecule. Since iodine dissolves in ethyl alcohol, molecules of ethyl alcohol probably contain covalent bonds ("like dissolves like").
13. **a.** Steps should include
(1) measure a mass of solute using a balance
(2) dissolve the solute in water
(3) add more water to dilute the solution to the required volume
(4) mix the solution thoroughly
(5) transfer the solution into a clean, dry, WHMIS-labelled container
b. Steps should include
(1) measure a volume of stock solution that will be diluted
(2) place stock solution in volumetric flask
(3) add water to flask to dilute solution
14. 17.14 mL
15. 4.5×10^{19} t
16. Safety— gloves, face shield, lab apron. The dilution must be done by adding the concentrated acid to water under a fume hood. Use a graduated cylinder to measure 72.9 mL of 12.0 mol/L $H_2SO_4(aq)$, and add this slowly, while stirring, to about 300 mL of distilled water in a 500 mL volumetric flask. (A purist would say this addition should be made in a beaker, because the heat generated will affect the calibration of the volumetric). Make the volume up to the calibration mark using a wash bottle of distilled water.
17. Answers should include: Flame tests—potassium gives a lavender colour flame; the flame due to calcium is reddish-orange. Chemical test—add sodium carbonate to precipitate $CaCO_3(s)$, or sodium sulfate to precipitate $CaSO_4(s)$. The potassium salts are soluble.
18. **a.** $Ba(NO_3)_2(aq) + Na_2S(aq) \rightarrow 2NaNO_3(aq) + BaS(s)$
b. 0.2 mol/L
19. Nitrates, phosphates, and insecticides; run-off from fields. Dissolved carbon dioxide and sulfur dioxide; acidic gases resulting from burning fossil fuels. Organic liquids from business sources, such as dry cleaning solvents.
20. The nitrogen in urea would act as a fertilizer to plants in rivers and lakes. When plants die they are decomposed by bacteria, which consume oxygen in the process. The removal of oxygen from river and lake water puts fish at risk.
21. **a.** Take a sample of the soil and place it in a flask. Add some distilled water and shake. Once the soil has settled, use universal pH paper to measure the pH of the water above the settled soil.
b. acidic
c. Apply ground limestone or wood ash (which contains calcium carbonate).
22. **a.** The acidic hydrogen atom is the one on the far right in the diagram. This H atom is highly polar due to the electronegativity of the oxygen atom attached to it. The C atom this oxygen atom is attached to, is made more polar by the attachment of another oxygen atom.
b. $CH_3CHOHCOOH(aq) + H_2O(\ell) \rightarrow$
$CH_3CHOHCOO^-(aq) + H_3O^+(aq)$
c. weak; it is not one of the strong acids; the acidic hydrogen atom is not as polar as the H atoms in the strong acids

23. The student will add less base than he thinks he is adding. Therefore, his calculated concentration of the acid will be lower than the actual concentration of the acid. The student should stand on a stepstool when the solution level in the burette is above his head. He should always read the burette level to the solution in it.

24. The reaction would generate noxious fumes and a lot of heat that would generate a dangerously hot or even boiling solution. Also, adding too much strong base would result in a solution of a strong base, which is equally hazardous.

25. 76.9 mL