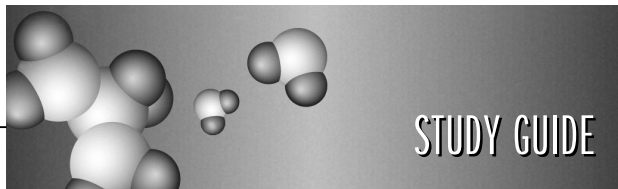


Name _____

Date _____

Class _____



1.1 The Puzzle of Matter

CHAPTER 1

Use with text pages 4 – 33

Use each of the terms just once to complete the following paragraphs.

- | | | |
|------------------|-------------|--------------|
| alloy | homogeneous | properties |
| aqueous solution | mass | qualitative |
| chemistry | matter | quantitative |
| heterogeneous | mixture | substances |

The science that deals with the study of the “stuff” that makes up the universe is (1) _____, although a better name for “stuff” is (2) _____. The measure of the amount of “stuff” in an object is the (3) _____ of the object. A description of a material using measurements is called a(n) (4) _____ observation, compared to a(n) (5) _____ observation, such as color or odor. The kinds of observations you can make about an object, such as its size, shape, color, odor, and tendency to react with other substances, are called the (6) _____ of the object.

Materials can be classified as (7) _____ if they have the same composition throughout, or (8) _____ if their composition varies from region to region. Examples of the former are elements and compounds, both of which are classified as (9) _____ because of their fixed compositions. An example of the latter is a combination of sand and salt that can be separated rather easily. This combination is called a(n) (10) _____. All homogeneous mixtures are (11) _____. A solution of two or more metals such as brass is a(n) (12) _____.

For each item in Column A, write the letter of the best matching item in Column B.

Column A

13. _____ 3000 grams
14. _____ property
15. _____ drawing of a molecule
16. _____ 3.48 meters
17. _____ green
18. _____ oxygen
19. _____ carbon dioxide
20. _____ sugar water
21. _____ brass
22. _____ boiling
23. _____ Sn
24. _____ CaCO_3

Column B

- a. scientific model
b. name of an element
c. chemical symbol of an element
d. aqueous solution
e. physical change
f. chemical formula
g. color, odor, weight, size, shape, hardness
h. name of a compound
i. qualitative observation
j. mass measurement
k. alloy
l. quantitative observation of length

Answer the following questions.

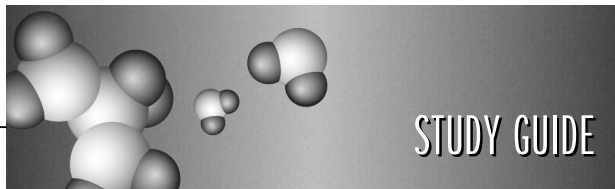
25. How do the submicroscopic and macroscopic views of matter differ?

26. How are the submicroscopic and macroscopic views related? _____

27. Suppose you make lemonade by dissolving a dry mix in water.

- a. What is the solute? _____
b. What is the solvent? _____
c. What is the aqueous solution? _____

28. The chemical formula for water is H_2O . What does this formula tell you about water?



1.2 Properties and Changes of Matter

CHAPTER 1

Use with text pages 34 – 44

In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- _____ 1. The freezing point of water is an example of a _____.
a. chemical change c. physical state
b. chemical property d. physical property
- _____ 2. The temperature most commonly taken as room temperature in scientific discussions is about _____.
a. 0°C c. 20°C
b. 100°C d. 37°C
- _____ 3. The name given to the change of state from gas to liquid is _____.
a. condensation c. freezing
b. evaporation d. cooling
- _____ 4. If a substance has an odor, it is _____.
a. a solid c. volatile
b. a liquid d. a pure substance
- _____ 5. The density of an object with a volume of 3.0 mL and a mass of 15.0 g is _____.
a. 45.0 g/mL c. 0.2 g/mL
b. 5.0 g/mL d. 12.0 g/mL
- _____ 6. The material with the lowest density among those listed is _____.
a. gold c. hydrogen gas
b. water d. gasoline
- _____ 7. An example of a chemical change is _____.
a. the boiling of water c. the rusting of iron
b. the evaporation of alcohol d. the breaking of glass
- _____ 8. Suppose six atoms of oxygen are present at the beginning of a chemical reaction. How many oxygen atoms are present when the reaction is over?
a. always six c. more than six
b. fewer than six d. no way of telling
- _____ 9. Materials like pewter are classified as a(n) _____.
a. aqueous solutions c. alloys
b. heterogeneous solutions d. compounds

For items 10–17, classify each of the following as a chemical or physical change.

10. wind eroding rocks _____
11. dead leaves decaying _____
12. rain puddle drying up _____
13. mixing flour and baking powder _____
14. gasoline evaporating _____
15. hydrogen peroxide decomposing _____
16. bread baking in an oven _____
17. instant tea dissolving in water _____
18. A 15 mL sample of a solid is weighed. The mass is recorded as 29 g. What is the density of the solid? _____

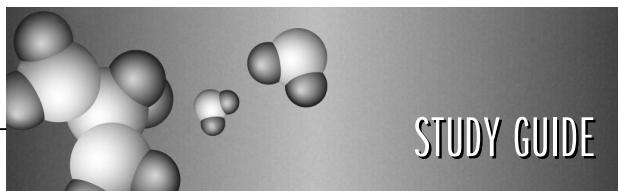
Identify each of the following changes as endothermic or exothermic.

19. decomposition of water _____
20. burning wood _____
21. photosynthesis _____
22. explosion of fireworks _____

Name _____

Date _____

Class _____



2.1 Atoms and Their Structure

CHAPTER 2

Use with text pages 52 – 68

Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 1. The ancient Greeks believed that matter is made of four basic elements: earth, air, fire, and oxygen.
- _____ 2. The man who proposed the first modern atomic theory of matter was Antoine Lavoisier.
- _____ 3. The law of definite proportions says that the composition of a compound is always the same.
- _____ 4. In science, a prediction that can be tested is known as a theory.
- _____ 5. The tiny negatively charged particles discovered by J.J. Thomson in 1897 were electrons.
- _____ 6. The number of protons in an atom is always equal to the number of neutrons.
- _____ 7. Isotopes differ from one another in their chemical properties.
- _____ 8. Rutherford's gold foil experiment led to the discovery of the portion of the atom known as the nucleus.
- _____ 9. The largest portion of an atom is taken up by the nucleus.
- _____ 10. The mass number of an atom is equal to the total number of protons and neutrons.
- _____ 11. Scientists use the symbol *g* to represent the atomic mass unit.
- _____ 12. An atomic mass unit is about equal to the mass of one proton.

Supply the missing information needed to complete the table below.

Element	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
13. sodium	11	23			
14. hydrogen		1	1		
15. neon			10	10	
16. fluorine				10	9
17. uranium	92			146	
18. aluminum				14	13

Answer the following questions.

19. Which two of the three statements that make up Dalton's atomic theory are no longer regarded as true as stated by Dalton? _____

20. Is it correct or incorrect in science to define a theory as someone's best guess about an aspect of nature? Explain your answer. _____

21. What evidence convinced J.J. Thomson that an atom consists of more than electrons? _____

22. What feature of Rutherford's gold foil experiment was most surprising to those who carried out the experiment? _____

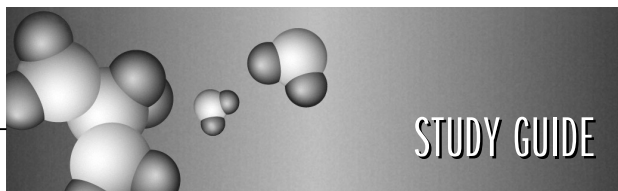
23. Could two isotopes of an element have different numbers of protons? Explain.

24. The element boron has two isotopes, with mass numbers of 10 and 11. The abundance of the two isotopes is 20 percent and 80 percent, respectively. Without making a mathematical calculation, estimate the atomic mass of boron.

Name _____

Date _____

Class _____



STUDY GUIDE

2.2 Electrons in Atoms

CHAPTER 2

Use with text pages 69 – 79

For each item in Column A, write the letter of the matching item in Column B.

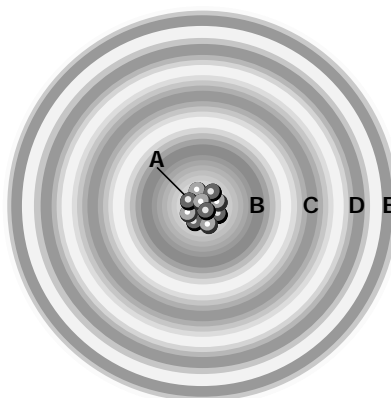
Column A

- _____ 1. electron
- _____ 2. empty space
- _____ 3. atomic nucleus
- _____ 4. Niels Bohr
- _____ 5. Lewis dot diagram
- _____ 6. hertz
- _____ 7. gamma rays
- _____ 8. radio waves
- _____ 9. emission spectrum
- _____ 10. energy level
- _____ 11. short wavelength
- _____ 12. valence electrons

Column B

- a. positively charged center of an atom
- b. electromagnetic energy with the highest frequency
- c. represents valence electrons
- d. negative particle found in an atom
- e. region around a nucleus where an electron can be found
- f. high-frequency wave
- g. most of an atom
- h. light released from excited atoms
- i. found in the outermost energy level
- j. suggested that electrons orbit the atom's nucleus
- k. unit of frequency of a wave
- l. electromagnetic radiation with the lowest energy

The diagram below shows a cross-sectional view of an atom. Answer the following two questions about this diagram.



13. What letter or letters on the diagram represents each of the following features of the atom?

- a. the nucleus: _____ c. the valence electrons: _____
b. energy levels: _____ d. electron cloud: _____

14. An emission spectrum containing 3 lines is obtained from an excited atom as shown in the diagram. For each line in Column A, write the letters of the matching transition shown in the diagram in Column B.

Column A

- _____ 1. line *x*, 400 nm
_____ 2. line *y*, 650 nm
_____ 3. line *z*, 750 nm

Column B

- a. C to B
b. D to C
c. D to B

Answer the following questions.

15. Write a Lewis dot diagram for each of the following elements.

- a. Na (11 electrons) d. Cl (17 electrons)
b. C (6 electrons) e. Ne (10 electrons)
c. Al (13 electrons) f. Sr (38 electrons)

16. Which form of electromagnetic radiation would you associate with each of the following descriptions?

- a. waves with energy just less than that of visible light _____
b. the form of electromagnetic energy with the shortest wavelength _____
c. electromagnetic energy with the longest wavelengths _____
d. electromagnetic energy with a frequency just less than that of gamma rays

17. Because electrons are traveling so fast, why do they not fly away from the atom?

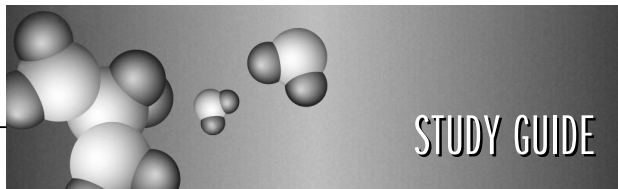
18. Chemistry teachers used to compare the structure of an atom to the structure of our solar system. In what way is that model correct? In what way is it not correct?

19. Why do you suppose that an energy level close to the nucleus can hold fewer electrons than one farther away from the nucleus? _____

Name _____

Date _____

Class _____



STUDY GUIDE

3.1 The Development of the Periodic Table

CHAPTER 3

Use with text pages 86 – 94

Use each of the terms below just once to complete the following passage.

triads

Döbereiner

atomic mass

noble gases

halogen

atomic numbers

periodicity

rows

Mendeleev

eka-aluminum

One of the earliest efforts to group the chemical elements in some kind of logical sequence was made by the German chemist (1) _____ in 1829. This effort was based on the fact that some elements have similar properties and can be placed in groups of three elements, called (2) _____. An example of such a group is the (3) _____ group, consisting of chlorine, bromine, and iodine. A more sophisticated method for grouping the elements was suggested in 1869 by the Russian chemist (4) _____. This method was based on the fact that, when the elements are arranged in order according to their (5) _____, their properties appear to repeat on a regular basis, a phenomenon known as (6) _____. In the first periodic table, elements with similar properties were placed together in (7) _____. Proof that this system of arranging the elements was correct came with the successful prediction of an element that had not yet been discovered, an element given the name of (8) _____. That element was actually discovered less than 20 years later. The primary way in which the modern periodic table differs from the original periodic table is that elements are now arranged according to their (9) _____. The rows in the modern periodic table begin with metals and end with (10) _____.

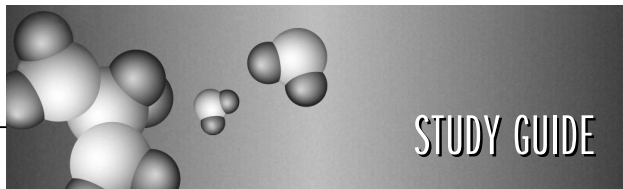
Circle the letter of the choice that is the best response or that best completes the statement.

11. In 1860, chemists could make which of the following statements about the known chemical elements?
- a. They all had identical properties. b. Some had similar properties.
c. They all had different properties. d. They could not be grouped.
12. Early scientists classified _____ as one of the coinage metals.
- a. lithium b. chlorine
c. copper d. iron
13. If the first and last elements in a triad have atomic masses of 7 and 39, respectively, then the atomic mass of the middle element is about _____.
- a. 23 b. 46
c. 15.3 d. 32
14. Mendeleev's first periodic table consisted of _____ groups.
- a. 3 b. 18
c. 8 d. 2
15. Which of the following is not a periodic phenomenon?
- a. rising of the sun b. appearance of Halley's comet
c. ocean tides d. growth of a sunflower
16. When moving from left to right across the modern periodic table, the atomic numbers of the elements _____.
- a. do not change b. increase by one
c. decrease by one d. are the same as the atomic masses
17. The modern periodic table is made up of _____ elements.
- a. 60 b. 111
c. 92 d. 25
18. In Döbereiner's halogen triad, the density, boiling point, and melting point of the elements _____ as atomic mass increases.
- a. stay the same b. increase
c. double d. decrease
19. In the modern periodic table, elements in the same _____ have similar properties.
- a. column b. period
c. row d. area
20. Döbereiner and Mendeleev both tried to relate the properties of elements to their _____.
- a. number of electrons b. atomic number
c. number of protons d. atomic mass

Name _____

Date _____

Class _____



3.2 Using the Periodic Table

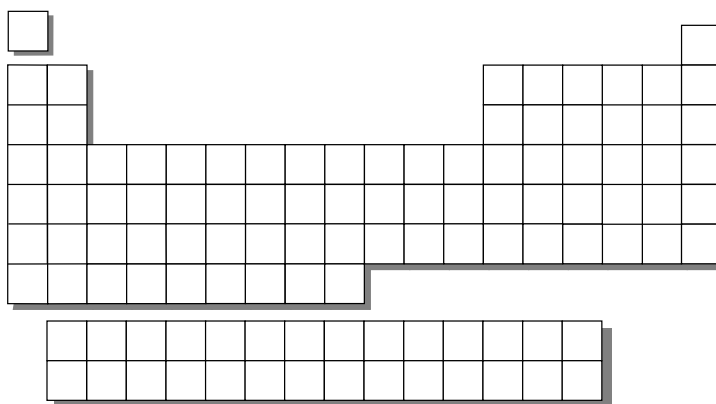
CHAPTER 3

Use with text pages 95 – 113

Select the group of elements that most closely fits each of the properties listed below. Write *M* for metals, *N* for nonmetals, and *O* for metalloids.

- _____ 1. luster
- _____ 2. tightly held valence electrons
- _____ 3. at the right side of the periodic table
- _____ 4. semiconductors
- _____ 5. good conductors of electricity
- _____ 6. usually quite brittle
- _____ 7. properties intermediate between metals and nonmetals
- _____ 8. a variety of colors
- _____ 9. three to eight valence electrons
- _____ 10. poor conductors of heat

Use the outline of the periodic table to complete the steps below. Each box represents an element.



- 11. Draw a straight blue line from the beginning to the end of any one period.
- 12. Draw a straight red line from the beginning to the end of any one group.
- 13. Place a V in any five boxes that you would expect to be metallic elements.
- 14. Place a W in any three boxes that you would expect to be nonmetallic elements.
- 15. Place an X in any box that you would expect to be a metalloid.
- 16. Place a Y in any box that you would expect to be a lanthanide element.
- 17. Place a Z in any box that you would expect to be an actinide element.

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For each item in Column A, write the letter of the matching item in Column B.

Column A

- _____ 18. hydrogen
 _____ 19. first period
 _____ 20. second period
 _____ 21. Group 2 elements
 _____ 22. magnesium
 _____ 23. oxygen
 _____ 24. noble gases
 _____ 25. transition elements
 _____ 26. metals
 _____ 27. nonmetals

Column B

- a. has six valence electrons
 b. are very unreactive
 c. have unpredictable properties
 d. has one more electron than sodium
 e. have loosely bound valence electrons
 f. has only one electron
 g. contains two elements
 h. have two valence electrons
 i. tend to gain electrons in chemical reactions
 j. contains eight elements

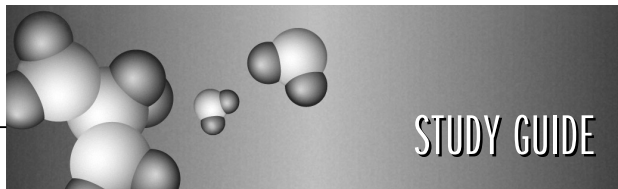
Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 28. A Lewis electron dot structure shows the total number of electrons in the atoms of an element.
 _____ 29. The electron dot structure for a metal is likely to have one, two, or three electrons.
 _____ 30. An element with an electron dot structure containing only one electron is likely to be found in Group 7 of the periodic table.
 _____ 31. Copper is a metal used in electrical circuits and in the manufacture of the alloy known as steel.
 _____ 32. Steel surfaces are sometimes plated with chromium to protect them from corrosion.
 _____ 33. Metalloids, such as silicon, are widely used in the manufacture of semiconductors.
 _____ 34. Pure silicon conducts an electrical current very well.
 _____ 35. The conductivity of silicon can be improved by doping it with other materials.

Name _____

Date _____

Class _____



4.1 The Variety of Compounds

CHAPTER 4

Use with text pages 120 – 129

Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

1. Sodium chloride is commonly known as table salt.

2. The most important source of sodium chloride is sea water.

3. Solid sodium chloride conducts an electrical current.

4. Elemental sodium is a silvery-white metal.

5. Sodium metal is normally stored under water.

6. Chlorine is one of the most reactive of all elements.

7. Carbon dioxide gas changes from a liquid to a solid at about -80°C .

8. The science of organic chemistry is concerned with compounds that contain silicon.

9. Oxygen makes up more than half of the Earth's air.

10. Hydrogen does not burn, but it is the gas that ordinarily supports the burning of other substances.

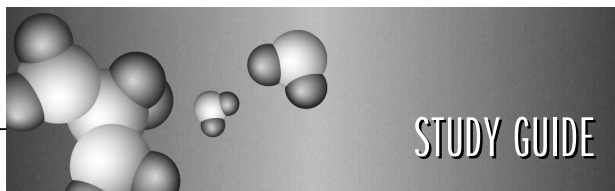
11. At -253°C , hydrogen changes from a gas to a liquid.

12. In general, compounds have properties that are the same as the properties of the elements of which they are composed.

From the following list of elements and compounds, select the one that best fits each of the descriptions below. Each item should be used more than once.

- | | | |
|--------------------|-------------------|-------------|
| a. sodium | d. carbon dioxide | g. hydrogen |
| b. chlorine | e. carbon | h. water |
| c. sodium chloride | f. oxygen | |

- _____ 13. Freezes at 0°C.
- _____ 14. Covers approximately 70 percent of Earth's surface.
- _____ 15. Is a colorless gas of low mass that is rarely found uncombined.
- _____ 16. Is a pale-green poisonous gas.
- _____ 17. Is the main constituent of charcoal.
- _____ 18. In its solid form, is called dry ice.
- _____ 19. Has crystals shaped like tiny cubes.
- _____ 20. Is produced when carbon burns in air.
- _____ 21. Is a metal that is soft enough to be cut with a knife.
- _____ 22. Is the most abundant element in Earth's crust.
- _____ 23. Along with silicon, makes up sand.
- _____ 24. Is often called the universal solvent.
- _____ 25. Is the most abundant element in the universe.
- _____ 26. Is a common source of heat.
- _____ 27. Is used as a disinfectant.
- _____ 28. Makes up about 0.03 percent of the atmosphere.
- _____ 29. Melts at close to the boiling point of water.
- _____ 30. Is the lightest of all elements.
- _____ 31. Can be obtained by the evaporation of sea water.
- _____ 32. Is located in the upper-left-hand corner of the periodic table.



4.2 How Elements Form Compounds

CHAPTER 4

Use with text pages 130 – 147

Use each of the terms below just once to complete the passage.

covalent bond	formula	ionic compound
covalent compound	interparticle force	molecule
crystal	ion	noble gas configuration
electrolyte	ionic bond	octet rule

When it reacts with a fluorine atom, a potassium atom loses one electron, forming a(n) (1) _____. Fluorine gains the electron and the charged atoms attract each other. This attraction is called a(n) (2) _____. The type of compound formed in this reaction, potassium fluoride, is called a(n) (3) _____. As a result of this reaction, the atoms of both potassium and fluorine have a stable set of outer electrons, an arrangement known as a(n) (4) _____. The statement that predicts that potassium will become stable by losing one electron and that fluorine will become stable by gaining one electron is the (5) _____. The regular, repeating arrangement in potassium fluoride is called a(n) (6) _____. Potassium fluoride, like other compounds of the same general type, is a(n) (7) _____, which is a substance that conducts an electric current when melted or dissolved in water.

In contrast to potassium fluoride, nitrogen monoxide is a(n) (8) _____ because nitrogen atoms in it share electrons with oxygen atoms in a kind of bond known as a(n) (9) _____. In contrast to potassium fluoride, the smallest unit of nitrogen monoxide is a(n) (10) _____. The force of attraction between particles, which is called (11) _____, is very different in these two compounds.

Both potassium fluoride and nitrogen monoxide are represented by a set of chemical symbols that represent their composition. The (12) _____ for potassium fluoride, for example, is KF, and for nitrogen monoxide, is NO.

In the spaces provided below, tell how the terms in each of the following pairs are different from each other.

13. ionic bond / covalent bond

14. ion / ionic bond

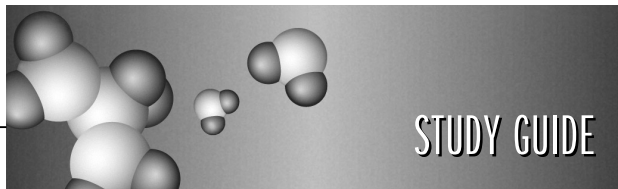
Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 15. Argon is a noble gas element.
- _____ 16. The stability of the noble gas elements other than helium is a result of their having eight valence electrons.
- _____ 17. Calcium has two valence electrons, and to become stable, it must gain two more electrons.
- _____ 18. An ion is an atom or group of atoms that is charged because it has lost or gained one or more protons.
- _____ 19. Ionically bonded compounds tend to be brittle.
- _____ 20. The formula C_3H_8 tells you that this compound contains three carbon atoms and eight hydrogen atoms.
- _____ 21. Covalent bonds involve electrons shared between two atoms.
- _____ 22. In the electron dot structure $H:\overset{\cdot\cdot}{\underset{\cdot\cdot}{Cl}}:$, the pair of dots between the H and the Cl means that the hydrogen and chlorine atoms transfer a pair of electrons between them.
- _____ 23. Two atoms cannot share more than one pair of electrons between them.
- _____ 24. In general, ionic compounds are more likely to dissolve in water than are covalent compounds.
- _____ 25. Ionic compounds conduct an electric current in the molten state.
- _____ 26. The forces of attraction between molecules tend to be strong.
- _____ 27. Sugar is an example of a covalent compound.

Name _____

Date _____

Class _____



5.1 Ionic Compounds

CHAPTER 5

Use with text pages 156 – 169

Write the word or phrase that best completes each of the following sentences or answers each question.

1. The charge on an ion is the same as the _____ of that ion.
2. A compound consisting of two elements is called a(n) _____.
3. A(n) _____ is an ion containing more than one element.
4. The “H₂O” portion of the formula BaCl₂ · H₂O indicates that this compound is a(n) _____.
5. Table salt sometimes becomes damp as it absorbs water from the air, indicating that table salt is a(n) _____ substance.
6. A(n) _____ substance takes up so much water from the air that it forms a liquid solution.
7. When hydrated copper sulfate is heated, it loses its water of hydration and becomes _____ copper sulfate.
8. A(n) _____ is the combination of chemical symbols that represents the simplest possible ratio of ions in a compound.
9. Magnesium bromide (MgBr₂) is one of the most abundant compounds found in sea water. From its formula, answer the following questions about this compound.
 - a. What is the oxidation number of each ion in the compound?

 - b. What would be the formula for the dihydrate of magnesium bromide?

 - c. How would you express the following, using chemical symbols:
“three formula units of magnesium bromide”?

10. In the chemical formula of a binary ionic compound, the last part is _____.

For each of the following formulas, write *true* if the formula is correct or change the formula to make it correct.

11. NaCl _____
12. AlF₃ _____
13. K₃PO₄ _____
14. CaNO₃ _____
15. BaCl₂, 2H₂O _____
16. FLi _____
17. MgSO₄ _____
18. Ca₂C₂H₃O₂ _____
19. K₂Cr₂O₇ _____
20. MgI _____

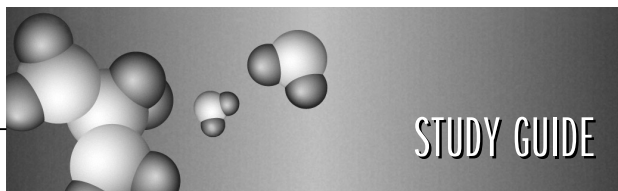
Write the correct name for each of the compounds in questions 11 – 20 above.

21. 11: _____
22. 12: _____
23. 13: _____
24. 14: _____
25. 15: _____
26. 16: _____
27. 17: _____
28. 18: _____
29. 19: _____
30. 20: _____

Name _____

Date _____

Class _____



STUDY GUIDE

5.2 Molecular Substances

CHAPTER 5

Use with text pages 170 – 185

For each of the properties listed below, write *I* for ionic compound or *M* for molecular substance to indicate which kind of substance is more likely to have that property.

- _____ 1. low melting point
- _____ 2. brittle
- _____ 3. conducts an electric current
- _____ 4. forms hydrates
- _____ 5. weak interparticle forces
- _____ 6. contains carbon
- _____ 7. contains bonds involving gain and loss of electrons
- _____ 8. can contain a polyatomic ion
- _____ 9. usually not very soluble in water
- _____ 10. contains bonds in which atoms share electrons

For each item in Column A, write the letter of the matching item in Column B.

Column A

- _____ 11. describes molecular substances
- _____ 12. example of allotropes
- _____ 13. describes organic compounds
- _____ 14. simplest hydrocarbon
- _____ 15. common inorganic compound

Column B

- a. methane
- b. water
- c. covalent bonding
- d. contain carbon
- e. oxygen and ozone

Fill in the missing name or formula for each compound.

	Name	Formula
16.		Ca(OH)_2
17.	sulfuric acid	
18.	ammonia	
19.		SO_2
20.		Na_2O
21.		CCl_4
22.	diphosphorus pentoxide	
23.	hydrochloric acid	
24.	octane	
25.		FeBr_3

In the space provided, tell why each formula or name of a compound is incorrect.

26. NaK _____

27. sulfur oxide _____

28. I_2Ca _____

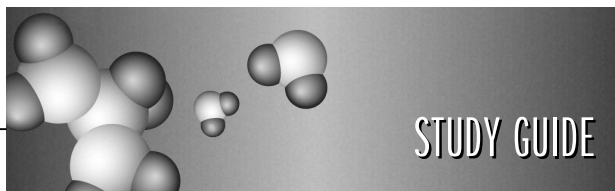
29. copper chloride _____

30. NH_4SO_4 _____

Name _____

Date _____

Class _____



STUDY GUIDE

6.1 Chemical Equations

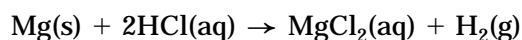
CHAPTER 6

Use with text pages 190 – 201

In each blank, write the word or phrase that best completes the following passage.

When a piece of magnesium metal is added to dilute hydrochloric acid, fizzing occurs and hydrogen gas is released from the mixture. The fizzing is evidence that a(n) (1) _____ has occurred between magnesium and hydrochloric acid. The name given to either magnesium or hydrochloric acid in this case is (2) _____, and the hydrogen gas that is released is called a(n) (3) _____ of the reaction. Some other indications that reactions have occurred might be change of color or (4) _____, or formation of a solid (5) _____. If a thermometer is placed into a mixture undergoing a reaction, you might observe that the temperature has gone up or down, indicating that (6) _____ was being released or absorbed. The shorthand form by which a reaction is represented is called a(n) (7) _____. In using this method of representation, you must satisfy the (8) _____, a principle that states that matter is neither created nor destroyed. In order to satisfy this principle, you normally select the proper numerical (9) _____ to indicate the number of units of each substance taking part in the chemical change.

In the space provided, express in words each of the numbered terms or symbols in the following chemical equation.



10. Mg _____

15. (aq) _____

11. (s) _____

16. \rightarrow _____

12. + _____

17. $\text{MgCl}_2\text{(aq)}$ _____

13. 2 _____

18. H_2 _____

14. HCl(aq) _____

19. (g) _____

Write a word equation and a balanced chemical equation for each of the reactions described below.

20. Aluminum metal burns in pure oxygen gas to produce solid aluminum oxide.

word equation: _____

chemical equation: _____

21. When solid mercury(II) oxide is heated, it breaks down to form liquid mercury and oxygen gas.

word equation: _____

chemical equation: _____

22. The addition of a solution of ammonium sulfate to a solution of lead(II) nitrate results in the formation of ammonium nitrate, which remains in solution, and lead(II) sulfate, which settles out of solution as a solid.

word equation: _____

chemical equation: _____

23. Copper metal and iron(II) nitrate in solution are formed when iron metal is added to a solution of copper (II) nitrate.

word equation: _____

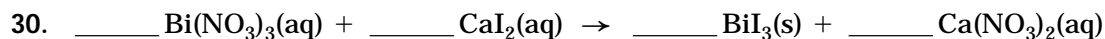
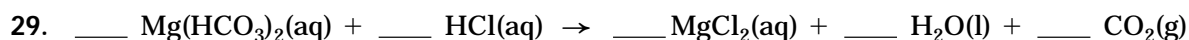
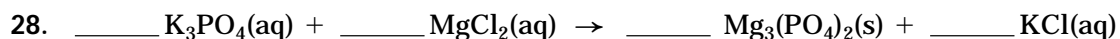
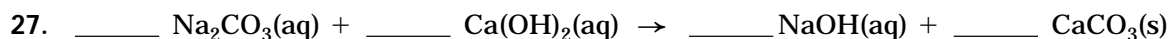
chemical equation: _____

24. Hydrogen sulfide gas reacts with pure oxygen gas to form water vapor and solid particles of sulfur.

word equation: _____

chemical equation: _____

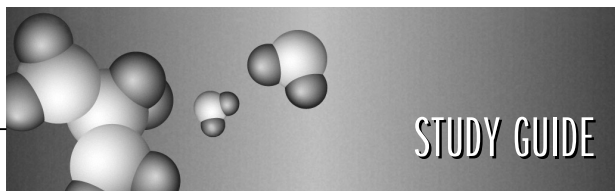
Balance each of the following chemical equations.



Name _____

Date _____

Class _____



STUDY GUIDE

6.2 Types of Reactions

CHAPTER 6

Use with text pages 202 – 209

Choose the correct symbol listed below to describe each of the chemical reactions listed in questions 1–15. Also write a balanced chemical equation for each reaction.

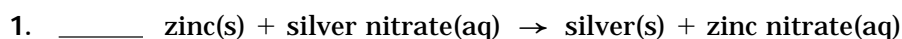
S = synthesis

D = decomposition

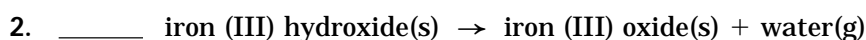
C = combustion

SD = single displacement

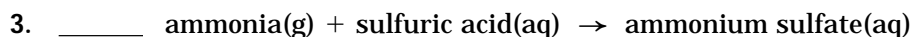
DD = double displacement



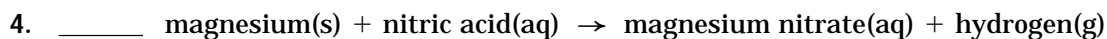
balanced equation: _____



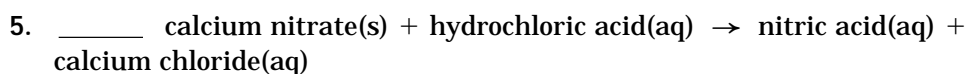
balanced equation: _____



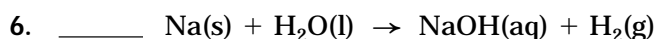
balanced equation: _____



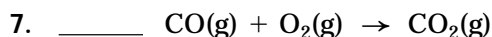
balanced equation: _____



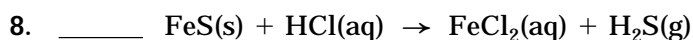
balanced equation: _____



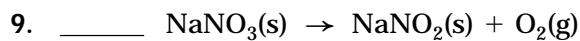
balanced equation: _____



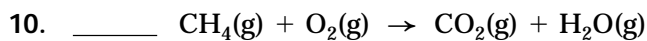
balanced equation: _____



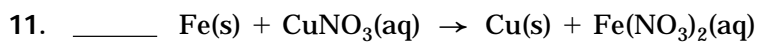
balanced equation: _____



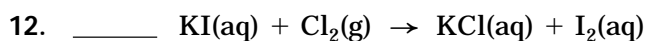
balanced equation: _____



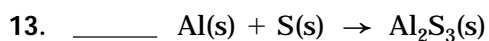
balanced equation: _____



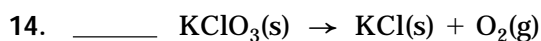
balanced equation: _____



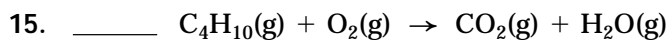
balanced equation: _____



balanced equation: _____



balanced equation: _____

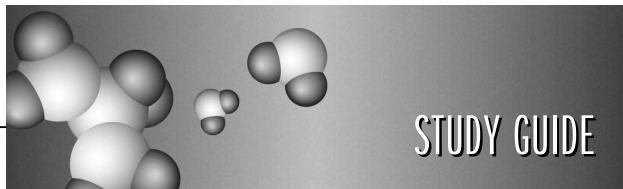


balanced equation: _____

Name _____

Date _____

Class _____



6.3 Nature of Reactions

CHAPTER 6

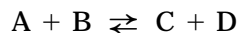
Use with text pages 210 – 223

In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- | | | |
|----------------------|----------------|----------------------|
| a. activation energy | d. enzyme | g. insoluble |
| b. catalyst | e. equilibrium | h. limiting reactant |
| c. concentration | f. inhibitor | i. soluble |

- _____ 1. The reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ can occur only when nitrogen molecules and hydrogen molecules have sufficient energy.
- _____ 2. The amounts of nitrogen, hydrogen, and ammonia in this reaction do not change.
- _____ 3. The rate of this reaction can be increased by adding a mixture of iron, potassium oxide, and aluminum oxide. The added materials are not permanently changed.
- _____ 4. The amount of ammonia produced by this reaction would be small if only a small amount of nitrogen were used.
- _____ 5. It is important to know the amount of each substance per unit volume.
- _____ 6. Some substances needed in the body of a living thing speed up reactions but are not used up.
- _____ 7. Some materials slow reactions down.
- _____ 8. Some materials cannot be made to dissolve in water.
- _____ 9. Some materials can dissolve in water.

For the chemical reaction described by the equation below, write three sentences that tell how to make the reaction occur more rapidly or more slowly. In each sentence, use a term from the list of nine words you worked with above.



More rapidly:

10. _____

11. _____

Name

Date

Class

12. _____

More slowly:

13. _____

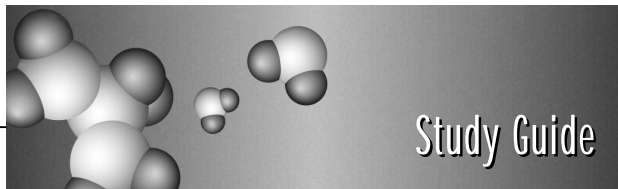
14. _____

15. _____

Name _____

Date _____

Class _____



7.1 Expanding the Theory of the Atom

CHAPTER 7

Use with text pages 229 – 242

In the spaces provided below, contrast terms in each pair.

1. Dalton atom and Rutherford atom _____

2. Energy level and sublevel _____

3. Frequency and wavelength: _____

4. Spectral lines from transitions between energy levels and spectral lines from transitions between sublevels _____

5. Maximum number of electrons in energy level 2 and in energy level 3

6. Electron cloud and orbital _____

7. *s* orbital and *p* orbital _____

8. Emission spectra and energy levels _____

9. Number of valence electrons and periodic table group number _____

10. Number of sublevels in energy level 2 and number of sublevels in energy level 3

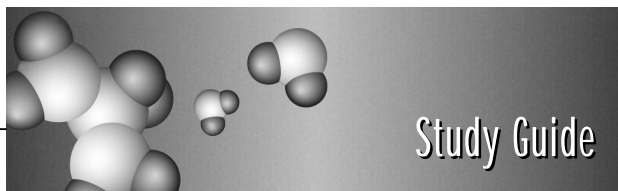
Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 11. Atoms of elements in the second row of the periodic table contain 1s, 2s, and 2p electrons.
- _____ 12. The Heisenberg uncertainty principle says that it is impossible to determine the position and the mass of an electron in an atom at the same time.
- _____ 13. The theory of the atom developed by Niels Bohr was valid for only one kind of atom, the helium atom.
- _____ 14. No *s* orbital can ever have more than two electrons in it.
- _____ 15. No *p* orbital can ever have more than six electrons in it.
- _____ 16. When an electron jumps from energy level 1*s* to energy level 2*s*, it absorbs the same amount of energy as it would release in falling from energy level 2*s* to energy level 1*s*.
- _____ 17. The maximum number of electrons in sublevel 4*f* is 32.
- _____ 18. An orbital is a space in which the probability of finding an electron is about 50 percent.
- _____ 19. Two orbitals can overlap with each other because they are filled primarily with empty space.
- _____ 20. The *s* orbitals in an atom are spherical in shape and are the same size.
- _____ 21. The farther an electron falls to its original level, the more toward the red end of the spectrum the color given off will be.
- _____ 22. Electrons in an atom are distributed in a way that makes an arrangement with the lowest energy.
- _____ 23. The overlap of orbitals results in the general shape of a sphere.
- _____ 24. Electrons with the least amount of energy are farthest from the nucleus.

Name _____

Date _____

Class _____



7.2 The Periodic Table and Atomic Structure

CHAPTER 7

Use with text pages 243 – 251

For each element in Column A, write the letter of the matching electron configuration in Column B.

Column A	Column B
1. _____ hydrogen	a. $[\text{He}]2s^32p^1$
2. _____ lithium	b. $[\text{He}]2s^22p^3$
3. _____ boron	c. $1s^22s^22p^6$
4. _____ carbon	d. $[\text{Ar}]4s^2$
5. _____ nitrogen	e. $[\text{Ne}]3s^23p^3$
6. _____ fluorine	f. $1s^1$
7. _____ neon	g. $[\text{Ne}]3s^23p^5$
8. _____ sodium	h. $1s^22s^22p^63s^23p^1$
9. _____ aluminum	i. $[\text{He}]2s^22p^2$
10. _____ phosphorus	j. $[\text{He}]2s^22p^5$
11. _____ chlorine	k. $[\text{He}]2s^1$
12. _____ calcium	l. $1s^22s^22p^63s^1$

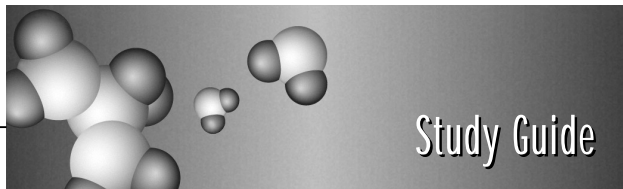
In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- _____ 13. In the electron configuration of an element, the first electron always goes into _____.
- a. the $1s$ orbital c. the nucleus
b. the $2s$ orbital d. an orbital that depends upon the element
- _____ 14. The correct order for filling orbitals in an atom is _____.
- a. $1s, 2s, 3s, 4s$ c. $1s, 2s, 3s, 2p$
b. $1s, 2s, 2p, 3s$ d. $1s, 1p, 2s, 2p$
- _____ 15. Groups 3 through 12 in the periodic table are designated as the _____.
- a. s region c. d region
b. p region d. f region

Name _____

Date _____

Class _____



8.1 Main Group Elements

CHAPTER 8

Use with text pages 258 - 281

From the following list of elements, select the one that best fits each of the descriptions below.

aluminum

iodine

phosphorus

boron

magnesium

potassium

calcium

nitrogen

sulfur

fluorine

oxygen

tin

- _____ 1. The most abundant element on Earth
- _____ 2. Needed for proper functioning of the thyroid gland
- _____ 3. Reacts violently with water
- _____ 4. Used to make alloys such as bronze, solder, and pewter
- _____ 5. A Group 15 element used in making matches
- _____ 6. An important element in bones and teeth
- _____ 7. Extracted from underground mines with hot water
- _____ 8. Its compounds help prevent tooth decay.
- _____ 9. Its atoms are found at the center of chlorophyll molecules.
- _____ 10. The most abundant metallic element on Earth
- _____ 11. Used in the production of water softeners and certain types of glass
- _____ 12. Obtained from liquid air by fractional distillation

Answer the following questions.

- 13. Why do elements in the same period have properties that are different from one another? _____

- 14. In general, how does atomic radius change across a period from left to right? Why?

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15. How does the size of a metallic atom change when it loses electrons to become a positive ion? _____
16. How does the size of a metallic atom change when it gains electrons to become a negative ion? _____
17. What type of configuration do elements tend toward as they react?

18. Why does boron not always continue the Period 2 pattern of reacting by losing electrons, as do beryllium and lithium? _____

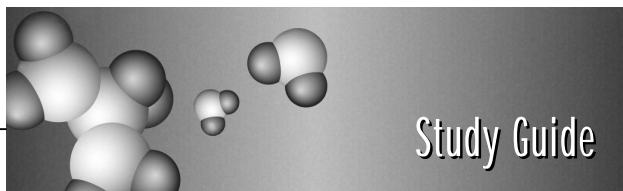
Compare each of the following pairs of elements, atoms, or ions by underlining the correct word in the pair.

19. Cesium is more/less active than sodium.
20. Sodium is more/less active than magnesium.
21. An aluminum ion has a larger/smaller radius than a magnesium ion.
22. An aluminum ion has a larger/smaller radius than a gallium ion.
23. Carbon has more/fewer valence electrons than does oxygen.
24. An atom of fluorine has a larger/smaller radius than does an atom of iodine.
25. Sulfur is more/less active than selenium.
26. Antimony is more/less metallic than is arsenic.
27. A halogen is more/less likely to occur in the free state than are members of the nitrogen family.
28. An atom of scandium is larger/smaller than an atom of titanium.

Name _____

Date _____

Class _____



8.2 Transition Elements

CHAPTER 8

Use with text pages 282 – 295

From the following list of elements, select the one that best fits each of the descriptions below.

americium
cerium
chromium
plutonium

cobalt
iron
mercury

silver
uranium
zinc

- _____ 1. Used as a power source in heart pacemakers
- _____ 2. The major component of steel
- _____ 3. The element most commonly used in making barometers
- _____ 4. Traditionally used in the manufacture of jewelry
- _____ 5. Frequently added to steel to make a harder, stronger product
- _____ 6. Used in the manufacture of ionizing smoke detectors
- _____ 7. Alloyed with copper to make brass
- _____ 8. Makes up about half of the alloy known as misch metal
- _____ 9. A naturally occurring element used as a nuclear fuel
- _____ 10. Like iron and nickel, is naturally magnetic

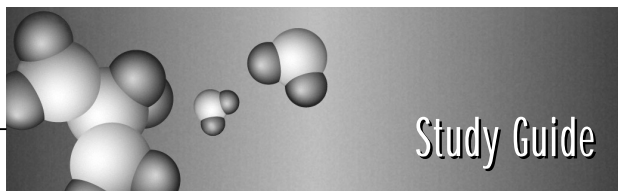
In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- _____ 11. A transition element known for its excellent ability to conduct an electric current is _____ .
 - a. iron
 - b. copper
 - c. uranium
 - d. silicon
- _____ 12. The element with the highest melting point is _____ .
 - a. iron
 - b. platinum
 - c. tungsten
 - d. iridium
- _____ 13. Which of the following elements has the lowest melting point?
 - a. sodium
 - b. copper
 - c. gallium
 - d. mercury

Name _____

Date _____

Class _____



9.1 Bonding of Atoms

CHAPTER 9

Use with text pages 302 – 314

For each item in Column A, write the letter of the matching item in Column B.

Column A

- _____ Ability to be drawn into a thin wire
- _____ A “sea of electrons”
- _____ Electrons shared unequally between two atoms
- _____ The ability of a substance to carry an electric current
- _____ The measure of an atom’s ability to attract electrons in a chemical bond
- _____ Ability to be hammered or rolled into a thin sheet
- _____ A decrease in the attraction exerted by an atomic nucleus on valence electrons
- _____ $\Delta EN = 0.3$
- _____ $\Delta EN = 3.0$
- _____ Symbolized by δ

Column B

- shielding effect
- partial charge
- ductility
- ionic bond
- nonpolar covalent bond
- polar covalent bond
- metallic bond
- malleability
- electronegativity
- conductivity

Use a table of electronegativities to calculate the value of ΔEN for each of the following pairs of atoms and predict the type of bond formed between them in each case.

11. lithium and sulfur

$\Delta EN =$ _____ Type of bond: _____

12. carbon and bromine

$\Delta EN =$ _____ Type of bond: _____

13. fluorine and silicon

$\Delta EN =$ _____ Type of bond: _____

14. magnesium and nitrogen

$\Delta EN =$ _____ Type of bond: _____

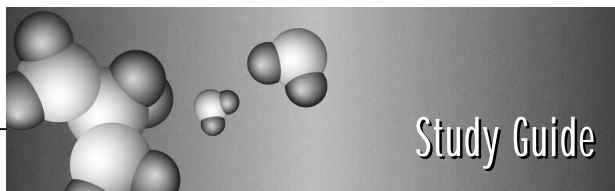
15. hydrogen and germanium

$\Delta EN =$ _____ Type of bond: _____

Name _____

Date _____

Class _____



9.2 Molecular Shape and Polarity

CHAPTER 9

Use with text pages 315 – 333

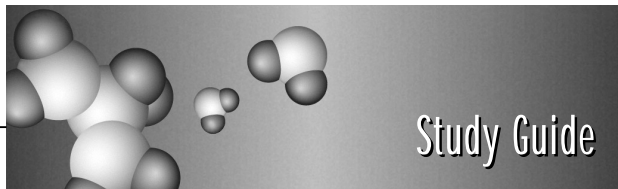
Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 1. The hydrogen molecule is linear.
- _____ 2. In a water molecule, four electron pairs are involved in covalent bonds.
- _____ 3. The approximately tetrahedral arrangement of electron pairs in a water molecule allows the electrons to be as far from each other as possible.
- _____ 4. The three-dimensional shape of a molecule is represented by a molecular formula.
- _____ 5. The angle between the two bonds in a water molecule is 109.5°.
- _____ 6. Triatomic molecules are always bent.
- _____ 7. A double bond is a covalent bond in which two electrons are shared between atoms.
- _____ 8. Four electron clouds surround the carbon atom in a carbon dioxide molecule.
- _____ 9. A methane molecule can be pictured as a tetrahedron.
- _____ 10. Methane is an example of an alkane, a compound whose molecules contain only carbon, hydrogen, and single bonds.
- _____ 11. Ethane and ethene have molecular shapes that are the same.
- _____ 12. The two carbons in ethyne are joined by double bonds.
- _____ 13. Ethyne is the simplest member of a family of organic compounds known as alkynes.
- _____ 14. A nonpolar molecule is one with a positive and a negative end.
- _____ 15. Another name for a polar molecule is a dipole.
- _____ 16. Polar molecules tend to be attracted to each other.
- _____ 17. The melting point of compounds that have polar molecules tends to be lower than the melting point of compounds that have non-polar molecules.

Name _____

Date _____

Class _____



10.1 Physical Behavior of Matter

CHAPTER 10

Use with text pages 340 – 347

From the following list of terms, select the one that best completes each sentence.

- ideal gas
- amorphous material
- kinetic theory
- Brownian motion
- plasma
- liquid
- solid
- gas
- crystal lattice
- liquid crystal
- ideal gas

The (1) _____ states that submicroscopic particles of matter are in constant motion. (2) _____, which is the constant, random movement of tiny chunks of matter, provides evidence for this model. In a(n) (3) _____, the particles are fixed in place, resulting in the formation of a highly organized structure known as a(n) (4) _____. The particles that make up a (5) _____, on the other hand, can flow, and matter in this state has an indefinite shape but a definite volume. A(n) (6) _____ has no definite volume or shape and is compressible. A(n) (7) _____ is one in which the particles undergo elastic collisions.

Some less familiar forms of matter also exist. In a(n) (8) _____, for example, particles may consist of rod-like molecules that have a rigid structure in only one or two dimensions. In comparison, a(n) (9) _____ appears to be solid, but has no well-organized structure. Finally, the most common form of matter in the universe is (10) _____, an ionized gas.

In the space provided, write the letter of the form of matter that best fits each description.

- solid (S)
- liquid (L)
- gas (G)
- plasma (P)
- liquid crystal (LC)
- amorphous material (AM)

- 11. _____ is used in calculator and watch displays.
- 12. _____ takes the shape of its container but has a definite volume.

13. _____ completely fills a container.
14. _____ is rigid, with a definite crystalline structure.
15. _____ retains an organized structure in one or two dimensions when melted.
16. _____ appears to be a solid, but has a haphazard structure.
17. _____ is formed at very high temperatures.
18. _____ is rare on Earth, but is in fluorescent lamps.
19. _____ flows and can be compressed to a smaller volume.
20. _____ maintains its volume but flows.

Answer the following questions.

21. Gases can be compressed easily. Liquids and solids can hardly be compressed at all. What statement about the particles that make up these states of matter can you make on the basis of these facts? _____

22. Define pressure, and explain how gases are able to exert pressure.

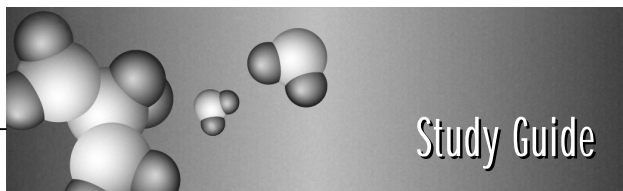
23. Use the kinetic model to describe the behavior of liquids. _____

24. Explain what plasmas are and how they are formed. _____

Name _____

Date _____

Class _____



10.2 Kinetic Energy and Changes of State

CHAPTER 10

Use with text pages 348 – 365

Explain how the terms in the following pairs are related to each other.

1. Kelvin and Celsius temperatures

2. temperature and the speed of particles

3. kinetic energy and mass of a particle, assuming constant speed

4. evaporation and sublimation

5. rate of diffusion and speed of particles

6. evaporation and boiling

7. vapor pressure and tendency to evaporate

8. interparticle forces and boiling point, assuming similar molecular mass

9. absolute zero and 0 K

10. atmospheric pressure and boiling point of a given liquid

Answer the following questions.

11. What is the value of each of the following temperature measurements on the Kelvin scale?

a. 78°C _____

b. -64.3°C _____

12. What is the value of each of the following temperature measurements on the Celsius scale?

a. 49 K _____

b. 432 K _____

13. Wet laundry hung outside on a cold day will dry even if the temperature is below freezing. How can you explain this fact? _____

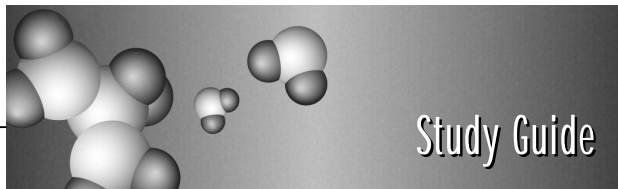
14. What is meant by the fact that the heat of vaporization of water is $2.26 \times 10^6 \text{ J/kg}$? Given that this quantity has the value $2.26 \times 18 \text{ J/kg}$ for water, how much energy will be needed to vaporize 175 g of water? _____

15. Describe the shape of a graph of temperature versus time for a solid being warmed, then melted to a liquid, that then is warmed to its boiling point and vaporized. Account for the shape. _____

Name _____

Date _____

Class _____



11.1 Gas Pressure

CHAPTER 11

Use with text pages 372 – 381

In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- _____ 1. At constant temperature and volume, the pressure exerted by a gas on the walls of a container is directly proportional to _____.
- a. the number of particles c. both the number of particles and mass
- b. the mass of the gas d. neither the number of particles nor the mass
- _____ 2. If the volume, number of particles and mass of a gas are not changed, the pressure of a gas is directly proportional to its _____.
- a. Celsius temperature c. Boyle temperature
- b. Kelvin temperature d. Fahrenheit temperature
- _____ 3. The barometer was invented by _____.
- a. Robert Boyle c. Evangelista Torricelli
- b. Jacques Charles d. Augustus Sphygmamus
- _____ 4. The SI unit for measuring pressure is the _____.
- a. atmosphere c. kilopascal
- b. psi d. pascal
- _____ 5. At sea level, atmospheric pressure is equal approximately to _____.
- a. 101.3 kPa c. 1000 atm
- b. 14.7 mm Hg d. 101.3 Pa
- _____ 6. The name of the device used to measure blood pressure is the _____.
- a. barometer c. sphygmomanometer
- b. pressure gauge d. pascal meter

Convert each of the measurements below, as directed.

7. Convert each of the following measurements to its equivalent in mm Hg.
- a. 1.2 atm _____
- b. 44.2 kPa _____
8. Convert each of the following measurements to its equivalent in atm.
- a. 785.2 mm Hg _____
- b. 132.1 kPa _____

9. Convert each of the following measurements to its equivalent in kPa.

a. 790.3 mm Hg _____

b. 1.442 atm _____

Answer the following questions.

10. If you push your finger into a balloon and then take your finger away, the “dent” in the balloon does not remain. Why? _____

11. Why is atmospheric pressure lower on the top of a mountain than it is at sea level?

12. List three ways to raise the air pressure inside a balloon. _____

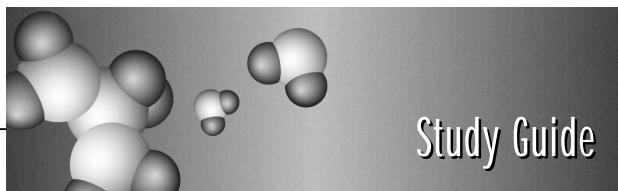
13. If you puncture a cylinder of carbon dioxide gas, why does the gas rush out instead of flowing or leaking out? _____

14. The valve on a tank of pressurized oxygen is opened and the oxygen rushes out. After oxygen gas stops flowing from a tank, some gas remains inside. Explain what determines how much gas is still in the tank. _____

Name _____

Date _____

Class _____



11.2 The Gas Laws

CHAPTER 11

Use with text pages 382 – 398

Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 1. The person who discovered the law expressing the quantitative relationship between the pressure and volume of a gas was Robert Boyle.
- _____ 2. The term direct proportion means that two related variables always increase together or always decrease together.
- _____ 3. The graph of a direct proportion is always a curved line.
- _____ 4. If the volume of a gas is 10 L at 1.5 atm, then the volume will be 5 L at 0.75 atm assuming the pressure does not change.
- _____ 5. Jacques Charles discovered the quantitative relationship between the temperature and volume of a gas while doing research on hot-air balloons.
- _____ 6. If the volume of a gas is 10 L at 100 K, then its volume will be 5 L at 200 K assuming the pressure does not change.
- _____ 7. To solve problems involving Charles' Law, one must always express the temperature in degrees Celsius.
- _____ 8. The term *standard temperature* means zero degrees kelvin.
- _____ 9. The volume of a gas whose temperature is doubled and whose pressure is doubled will increase four times.
- _____ 10. According to the law of combining gas volumes, the ratio in which gases combine or decompose is always a small whole number ratio.
- _____ 11. According to Avogadro's principle, equal volumes of gases at the same temperature and pressure contain equal numbers of particles.
- _____ 12. In the decomposition of water, the ratio of volumes of hydrogen to oxygen produced is 1:1.

Solve the following problems.

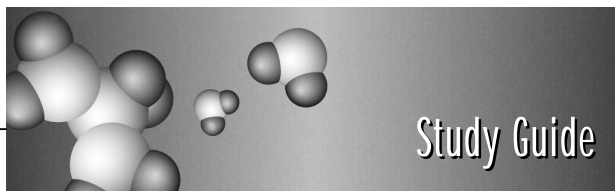
13. A balloon contains 750.0 mL of hydrogen gas at a pressure of 700.0 mm Hg. What will be the volume of that gas if the pressure is increased to 800.0 mm Hg?

14. A sample of 220.0 mL of neon gas is enclosed in a cylinder with a piston that can move up and down with changes in temperature. If the gas is at 10.0°C , and if the pressure on the gas does not change, what will be the volume of the gas if the temperature is increased by 90.0°C ? _____
15. A scientist wishes to compress 50.0 mL of ammonia gas at normal atmospheric pressure to a volume of 1.0 mL. What pressure in atmospheres is necessary?

16. In a study on carbon dioxide gas, a scientist begins with 30.0 mL of the gas at a temperature of 14°C . She lowers the temperature of the gas until its volume is reduced to 22.0 mL. The pressure on the gas is not changed during the experiment. To what temperature in $^{\circ}\text{C}$ was the gas lowered to reach this volume?

17. A sample of xenon gas occupying 10.0 mL is collected in a laboratory at 27°C and 1.25 atm. What is the volume of the gas at standard conditions?

18. A researcher is interested in finding the effects of reducing both the pressure and the temperature on a sample of uranium hexafluoride gas. If the gas occupies 6.6 L at standard conditions, what will its volume be at a -23°C and 150 mm Hg? _____
19. In a Boyle's law experiment, a student prepared a sample of oxygen gas on Monday and returned to study its properties on Tuesday. The volume of the gas was 75.0 mL on Monday and 70.0 mL on Tuesday. The temperature was 27°C on Monday and 25°C on Tuesday. On Tuesday, the student recorded the pressure on the gas as 4.4 atm, but forgot to record the pressure on the gas on Monday. What was that pressure in atmospheres? _____
20. A student studying the effects of cold temperatures on helium gas subjects a 10.0 mL sample of the gas originally at 20.0°C and 760 mm Hg to severe cooling. If the pressure on the gas rises to 800.0 mm Hg during the experiment, and the final volume of the gas is 1.0 mL, to what value in $^{\circ}\text{C}$ was the temperature reduced? _____



12.1 Counting Particles of Matter

CHAPTER 12

Use with text pages 404 – 413

In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- _____ 1. Which of the following quantities is *not* a mole of the substance?
- 15.999 g oxygen gas
 - 12.011 g carbon
 - 28.014 g nitrogen gas
 - 4.003 g helium gas
- _____ 2. The mass of a mole of zinc equals the atomic mass of zinc expressed in _____.
- mol
 - kilograms
 - grams
 - number of particles
- _____ 3. The number of atoms in one mole of sodium metal is _____.
- 23
 - 6.02×10^{23}
 - $23 \div (6.02 \times 10^{23})$
 - $(6.02 \times 10^{23}) \div 23$
- _____ 4. The term molecular mass is used only for compounds _____.
- containing carbon
 - consisting of ions
 - that are covalent
 - that do not ionize in water
- _____ 5. A sample of water whose mass is 18 g represents what quantity of water?
- formula mass
 - molar mass
 - atomic weight
 - atomic mass
- _____ 6. The formula mass for aluminum chloride is _____.
- 62.45 u
 - 6.02×10^{23} u
 - 97.9 u
 - 133.34 u
- _____ 7. Compared to the number of molecules in a mole of hydrogen gas (H_2), the number of molecules in a mole of water (H_2O) is _____.
- twice as many
 - the same
 - half as many
 - three times as many

- _____ 8. The mass of an average ruthenium atom is 8.4 times as large as the mass of an average carbon atom. What is the mass of one mole of ruthenium?
- 8.4 g
 - $8.4 \times 6.02 \times 10^{23}$ g
 - 100.9 g
 - $8.4 \times 12 \times 6.02 \times 10^{23}$ g
- _____ 9. The mass of one molecule of ammonia is its _____.
- molecular weight
 - molecular mass
 - formula mass
 - formula unit

Solve the following problems.

10. Find the number of atoms in each of the following measurements.
- 10.0 g of Al _____
 - 25.9 g of F _____
11. Find the molar mass of each of the following compounds.
- N_2O_5 _____
 - ammonium carbonate _____
12. Express each of the following in grams.
- 0.147 mol of NH_4NO_3 _____
 - 2.60 mol of lithium bromide _____
13. Convert each of the following to its equivalent in moles.
- 17.1 g of H_2S _____
 - 2.50 g of copper(II) chloride _____

Answer the following questions.

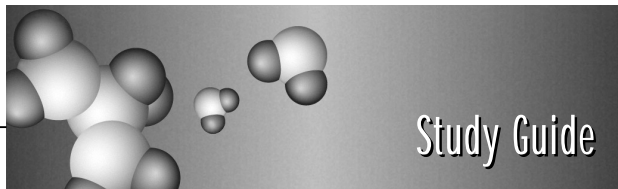
14. How many nitrogen atoms are in $\frac{1}{4}$ mole of nitrogen gas?

15. How many moles of oxygen atoms are in $\frac{1}{2}$ mole of $NaIO_4$?

Name _____

Date _____

Class _____



12.2 Using Moles

CHAPTER 12

Use with text pages 414 – 429

In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- _____ 1. In the equation $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, the number of molecules of water formed for every one molecule of hydrogen that reacts is _____.
- one
 - two
 - four
 - one-half
- _____ 2. In the equation in question 1, the number of molecules of oxygen needed to form 0.8 mol of water is _____.
- 0.4 mol
 - 0.8 mol
 - 0.2 mol
 - 1.6 mol
- _____ 3. The molar volume of a gas at standard conditions is _____.
- 1 L
 - 6.02×10^{23} L
 - 22.4 L
 - dependent on the gas
- _____ 4. In the equation that represents the ideal gas law, the only constant is _____.
- P
 - V
 - n
 - R
- _____ 5. In a chemical reaction in which 35.8 g of product is predicted and only 13.2 g obtained, the percent yield is _____.
- $35.8 \text{ g}/13.2 \text{ g} \times 100\%$
 - $13.2 \text{ g}/35.8 \text{ g} \times 100\%$
 - $35.8 \text{ g}/(35.8 \text{ g} + 13.2 \text{ g}) \times 100\%$
 - $(35.8 \text{ g} + 13.2 \text{ g})/35.8 \text{ g} \times 100\%$
- _____ 6. In the compound FeCl_3 , the percent by mass of chlorine in the compound is _____.
- mass of Cl/mass of $\text{FeCl}_3 \times 100\%$
 - mass of FeCl_3 /mass of Cl $\times 100\%$
 - $3 \times$ mass of Cl/mass of $\text{FeCl}_3 \times 100\%$
 - mass of $\text{FeCl}_3/3 \times$ mass of Cl $\times 100\%$

- _____ 7. If the empirical formula of a compound is HF and analysis shows its molar mass is 40 g, its molecular formula would be _____.
- HF
 - H_2F_2
 - H_3F_3
 - $\text{H}_{40}\text{F}_{40}$

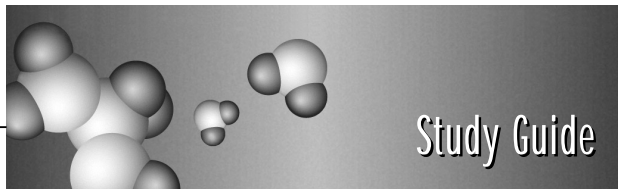
Solve the following problems.

8. Zinc metal reacts with hydrochloric acid according to the following equation:
 $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$. How many grams of hydrogen gas can be obtained if 50.0 g of zinc react completely?
- _____
9. What mass of magnesium chloride is needed to react with 145 g of silver nitrate according to the following reaction: $\text{MgCl}_2(\text{aq}) + 2\text{AgNO}_3(\text{aq}) \rightarrow \text{Mg}(\text{NO}_3)_2(\text{aq}) + 2\text{AgCl}$?
- _____
10. When potassium chlorate is heated, oxygen is produced according to the following reaction: $2\text{KClO}_3 + \text{heat} \rightarrow 2\text{KCl} + 3\text{O}_2$. How many liters of oxygen gas at STP can be obtained if 500.0 g of potassium chlorate react completely?
- _____
11. How many liters of hydrogen gas at STP are needed to react completely with 50.0 L of chlorine gas at STP in the following reaction: $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$? How many liters of hydrogen chloride gas at STP are produced in the reaction?
- _____
12. What volume will 2.39 mol of carbon dioxide gas occupy at of 32.5°C and 105 kPa?
- _____
13. What pressure will be needed to compress 5.00 mol of nitrogen gas at 23°C to a volume of 1.5 L?
- _____
14. A student heats 100.0 g of copper(II) sulfate pentahydrate and obtains 58.2 g of anhydrous copper(II) sulfate. What percent yield was obtained in this experiment? $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} + \text{heat} \rightarrow \text{CuSO}_4 + 5\text{H}_2\text{O}$
- _____
15. Find the percent composition by mass of phosphoric acid, H_3PO_4 . _____
- _____
16. Find the empirical formula for a compound that has the following percentage composition by mass: 10.04 percent carbon; 0.84 percent hydrogen; 89.12 percent chlorine.
- _____

Name _____

Date _____

Class _____



13.1 Uniquely Water

CHAPTER 13

Use with text pages 436 – 450

In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- _____ 1. The states in which water occurs on Earth are _____.
- a. liquid and gaseous only c. solid and gaseous only
b. liquid and solid only d. solid, liquid, and gaseous
- _____ 2. The physical property of ice that allows aquatic organisms to survive over the winter is its _____.
- a. high melting point c. low density
b. low boiling point d. high mass
- _____ 3. The water molecule is _____.
- a. polar and has polar bonds c. nonpolar and has polar bonds
b. polar and has nonpolar bonds d. nonpolar and has nonpolar bonds
- _____ 4. In comparison with covalent bonds, hydrogen bonds are _____.
- a. much stronger c. only slightly stronger
b. of about the same strength d. much weaker
- _____ 5. A compound besides water that would be expected to form hydrogen bonds is _____.
- a. CH₃CH₂CH₂CH₃ c. CH₂=CHCH₃
b. CH₃CHOHCH₂CH₃ d. CCl₄
- _____ 6. During cooling from 50°C to 0°C, the density of water _____.
- a. first increases, then decreases c. increases
b. first decreases, then increases d. decreases
- _____ 7. The skinlike quality of the surface of water is caused by water's _____.
- a. density c. covalent bonding
b. boiling point d. surface tension
- _____ 8. Compared to its height in a wide-diameter tube, the height of a column of water in a narrow-diameter tube will be _____.
- a. greater c. the same
b. less d. unpredictable

- _____ 9. Compared to the specific heat of most other materials, the specific heat of water is _____.
- a. about the same c. much higher
b. much lower d. higher than some, lower than others
- _____ 10. Which of the following processes is endothermic?
- a. cooling c. freezing
b. vaporization d. condensation
- _____ 11. Water seldom appears in a pure state on Earth because it tends to _____.
- a. evaporate c. dissolve other substances
b. change state so easily d. react with other substances

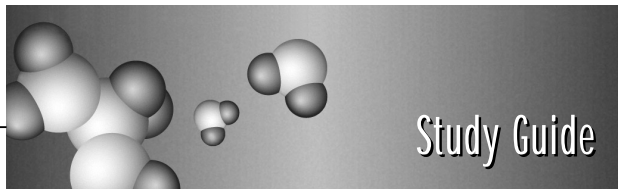
Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 12. Most of the physical properties of water are similar to the properties of other molecules of about the same size.
- _____ 13. The water molecule has two pairs of unbonded electrons.
- _____ 14. Water has its maximum density at a temperature of about 0°C.
- _____ 15. If lakes were made of ethanol rather than water, they would freeze from the bottom up.
- _____ 16. The hydrogen end of a water molecule tends to have a partial negative charge.
- _____ 17. The force of attraction between two different molecules is known as an intermolecular force.
- _____ 18. One would expect the hydrogen bonds in ethanol to be stronger than those in water.
- _____ 19. The hydrogen bonds in a protein molecule will break more easily than the covalent bonds in the molecule.
- _____ 20. The tendency to form liquid drops is less in water than it is in ethanol.
- _____ 21. The stronger the hydrogen bonds in a liquid, the more curved will be its meniscus in a capillary tube.
- _____ 22. The amount of heat needed to raise the temperature of 1 g of iron by 1°C is greater than the amount of heat needed to raise the temperature of 1 g of water by the same amount.
- _____ 23. Substances made of nonpolar molecules are likely to dissolve in water.

Name _____

Date _____

Class _____



13.2 Solutions and Their Properties

CHAPTER 13

Use with text pages 451 – 473

Write the word or phrase that best completes each of the following sentences.

1. When a crystal of sodium chloride is dissolved in water, the oxygen ends of the water molecules are attracted to the _____ of the sodium chloride.
2. The usual method for representing a chloride ion in water solution is as _____.
3. Sugar can be dissolved in water because of the presence of bonds between _____ in its molecules.
4. An equation that represents the dissolving of sucrose (sugar) in water is _____.
5. Because sugar does not conduct an electric current in water solution, it is described as a(n) _____.
6. In general, nonpolar substances do not dissolve in water, but they do dissolve in _____.
7. _____ refers to the relative amount of solute dissolved in a solvent.
8. If the maximum amount of a compound that can dissolve in 100 mL of water is 22 g, then a solution of 20 g of that substance in 100 mL of water is said to be _____.
9. A solution of 25 g of the substance described in question 8 in 100 mL of water would be called _____.
10. In most cases, one way to increase the solubility of a solid substance in water is to increase the _____ of the solution.
11. A more precise way for describing the concentration of a solution than to describe it as concentrated or dilute is to state its _____.
12. To make 5 L of 1M aluminum nitrate ($\text{Al}(\text{NO}_3)_3$), you would have to dissolve _____ of the solute in 5.0 L of solution.
13. 400 mL of a 0.1M barium nitrate ($\text{Ba}(\text{NO}_3)_2$) solution consists of _____ of solute dissolved in 400 mL of solution.

14. One mole of the compound aluminum nitrate lowers the freezing point of a solution by _____ times as much as one mole of sucrose.
15. Osmosis is the process by which molecules move through a _____.
16. The major difference between a solution and a colloid is _____.

Explain how the two terms in each of the following pairs differ from each other.

17. $C_6H_{12}O_6(s)$ and $C_6H_{12}O_6(aq)$ _____

18. concentrated and dilute _____

19. saturated and unsaturated _____

20. endothermic and exothermic _____

21. molarity and moles per liter of solvent _____

22. value of freezing point of water and value of freezing point of an aqueous solution

23. freezing point of $1M NaCl$ and freezing point of $1M C_6H_{12}O_6$ _____

24. solute and solvent _____

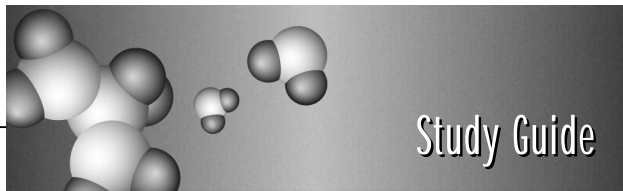
25. osmosis and reverse osmosis _____

26. colloid and solution _____

Name _____

Date _____

Class _____



14.1 Acids and Bases

CHAPTER 14

Use with text pages 480 – 496

For each item in Column A, write the letter of the best matching item in Column B.

Column A

- _____ 1. H_3PO_4
- _____ 2. indicator
- _____ 3. H_3O^+
- _____ 4. NH_3
- _____ 5. Na_2O
- _____ 6. sour taste
- _____ 7. slippery feel
- _____ 8. NaOH
- _____ 9. one acidic hydrogen
- _____ 10. $\text{HC}_2\text{H}_3\text{O}_2$
- _____ 11. H_2SO_4
- _____ 12. SO_3

Column B

- a. property of acid
- b. property of base
- c. litmus
- d. #1 industrial chemical
- e. polyprotic acid
- f. strong base
- g. weak base
- h. weak acid
- i. acidic anhydride
- j. basic anhydride
- k. $\text{HC}_2\text{H}_3\text{O}_2$
- l. hydronium ion

Use each of the terms just once to complete the following paragraph.

acid
acidic anhydride
acidic hydrogen
active metals

base
basic anhydride
carbonates

litmus
 H_3O^+
 OH^-

One way of detecting the presence of a(n) (13) _____ is by its bitter taste. This compound can also be recognized because it changes the color of (14) _____ from pink to blue. When placed in water, this type of compound releases (15) _____. Calcium hydroxide, a familiar example of this kind of compound, can be formed by adding calcium oxide to water, which makes calcium oxide a(n) (16) _____. In comparison, the addition of dinitrogen pentoxide (N_2O_5) to water results in the formation of HNO_3 , a(n) (17) _____. In this case, dinitrogen pentoxide is called a(n)

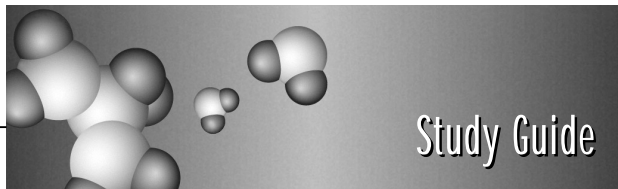
- (18) _____ . When HNO_3 is present in a water solution, it releases a(n) (19) _____ that combines with water to form a(n) (20) _____ . A typical reaction for HNO_3 involves the release of carbon dioxide gas from (21) _____ or the release of hydrogen from (22) _____ .

Write a chemical equation that illustrates each of the following points.

23. HBr is an acid. _____
24. Boron hydride (BH_3) is a base. _____
25. Any acid will react with calcium metal. _____
26. Any ionic base dissociates in water. _____
27. Any metal oxide (MO) is a basic anhydride. _____
28. Rain is naturally acidic. _____

Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 29. Drain and oven cleaners work because they are acids.
- _____ 30. Litmus is red in acidic solution.
- _____ 31. Acids react with metals that are more active than hydrogen to form a compound of the metal and hydrogen gas.
- _____ 32. NaOH is a base because it produces hydronium ions when it dissolves in water.
- _____ 33. Aqueous solutions of bases conduct electricity because a base in water produces ions.



14.2 Strengths of Acids and Bases

CHAPTER 14

Use with text pages 497 – 508

In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- _____ 1. A strong acid, compared to a weak acid, _____ .
a. has a stronger taste c. ionizes completely
b. contains only acidic hydrogens d. contains only two elements
- _____ 2. The only weak acid listed below is _____ .
a. HCl c. H_2SO_4
b. HClO_4 d. H_2S
- _____ 3. The only weak base listed below is _____ .
a. NaOH c. $\text{Ba}(\text{OH})_2$
b. HgOH d. $\text{Sr}(\text{OH})_2$
- _____ 4. The fraction of acetic acid molecules ionized in solution is about _____ .
a. 1% c. 50%
b. 25% d. 100%
- _____ 5. The acid most commonly found in soft drinks is _____ .
a. hydrochloric acid c. carbonic acid
b. carbolic acid d. sulfuric acid
- _____ 6. The acids and bases found in DNA and proteins are _____ .
a. weak acids and weak bases c. strong acids and weak bases
b. weak acids and strong bases d. strong acids and strong bases
- _____ 7. The pH of pure water is _____ .
a. 0 c. 7
b. 1 d. 14
- _____ 8. The pH most likely to be measured for a strong base is _____ .
a. 0 c. 8
b. 5 d. 13
- _____ 9. When hydrochloric acid is added drop-by-drop to a solution of sodium hydroxide, the pH of the mixture _____ .
a. stays the same c. increases
b. decreases d. changes in an irregular way

- _____ 10. A water solution of a strong acid is _____ .
- a. always concentrated
 - b. never concentrated
 - c. sometimes concentrated, sometimes not
 - d. always dilute
- _____ 11. In going from pH 4 to pH 5, the concentration of hydrogen ions _____ .
- a. increases by a factor of 10
 - b. decreases by a factor of 10
 - c. doubles
 - d. is cut in half
- _____ 12. All acid-base indicators _____ .
- a. change color at different pH values
 - b. have more than one endpoint
 - c. change from pink to blue or blue to pink
 - d. are made from natural products

In the space provided, explain how the meanings of the two terms in each pair differ from each other. For some terms, there may be no difference.

13. pH and hydrogen ion concentration _____

14. strong acid and weak acid _____

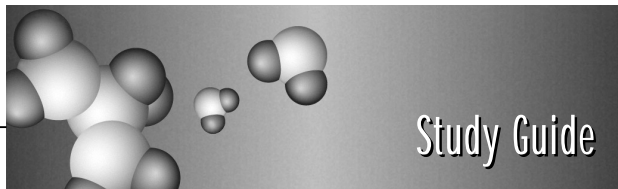
15. hydronium ion and hydrogen ion _____

16. sodium hydroxide and aluminum hydroxide _____

17. pH = 5 and pH = 9 _____

18. litmus and acid/base indicator _____
19. pH = 7 and neutral solution _____
20. pH meter and indicator paper _____
21. concentration and strength _____

22. pH of milk and pH of vinegar _____



15.1 Acid and Base Reactions

CHAPTER 15

Use with text pages 516 – 530

From the following list of terms, select the one that best completes each statement. Use each term only once.

acid-base indicator	base	salt	NH_4^+
HCN	H_2O	neutralization reaction	
H_2SO_4	Na^+	H_3O^+	

- BaCl_2 , $\text{Mg}(\text{NO}_3)_2$, and Na_2SO_4 are all examples of a(n) _____.
- Litmus paper is an example of a(n) _____.
- Because it breaks down completely into positive and negative ions in solution, _____ is considered a strong acid.
- Because it is usually a spectator ion, the chemical symbol _____ seldom appears in net ionic equations.
- Because it does not noticeably dissociate or ionize, the formula _____ often appears in net ionic equations.
- Because it represents a weak acid, the chemical formula _____ would be expected to appear in a net ionic equation.
- When ammonia occurs in water solution, it forms the ion _____.
- A substance that accepts hydrogen ions is a(n) _____.
- The addition of an acid to water can be considered an acid-base reaction because water molecules accept hydrogen ions to become _____ ions.
- By adding an acid to a base, you can observe a(n) _____.

Answer each of the following questions.

- Does the reaction between equal amounts of an acid and a base always produce a neutral solution? Why or why not? _____

- What color would you expect to see in a piece of litmus paper to which the following have been added?
 - 1 drop of $\text{Ca}(\text{OH})_2$ _____
 - 1 drop of HI _____
 - 1 drop of KNO_3 _____

13. What would the pH be of a solution formed by mixing the following?
- equivalent amounts of KOH and H_2SO_4 _____
 - equivalent amounts of KOH and HCN _____
 - equivalent amounts of NH_3 and H_2SO_4 _____

Write the overall, ionic, and net ionic equations for each of the following reactions.

14. *Reaction A:* nitric acid plus barium hydroxide yields water plus barium nitrate.

15. *Reaction B:* perchloric acid plus aluminum hydroxide yields water plus aluminum perchlorate

16. *Reaction C:* hydrocyanic acid (HCN) plus calcium hydroxide yields calcium cyanide plus water

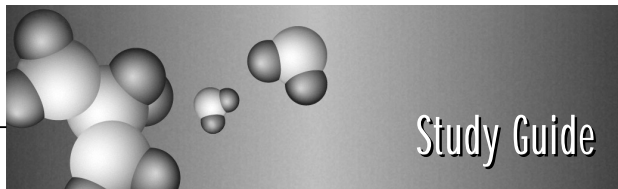
For each of the above reactions, tell whether you would expect the pH of the final solution to be: a: 7.0; b: greater than 7.0; c: less than 7.0.

17. Reaction A _____
18. Reaction B _____
19. Reaction C _____

Name _____

Date _____

Class _____



15.2 Applications of Acid-Base Reactions

CHAPTER 15

Use with text pages 531 – 546

Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 1. The normal pH of blood is about 2.5.
- _____ 2. A buffer is a solution that resists changes in pH when moderate amounts of acid and base are added.
- _____ 3. The amount of carbon dioxide in the blood buffer system is controlled by the lungs.
- _____ 4. The natural buffering that occurs in some lakes is a result of carbon dioxide in the air.
- _____ 5. The normal pH of stomach acid is about 6.5.
- _____ 6. Antacids are generally hydroxides or carbonates.
- _____ 7. In doing a titration, both of the solutions used must be a standard solution.
- _____ 8. To improve the results obtained in a titration, you should always repeat that titration several times.
- _____ 9. In most acid-base reactions, the oxidation numbers of the elements involved do change.

Complete each of the following statements.

10. The other half of a buffer system containing sodium citrate would be _____.
11. When hydroxide ions are added to blood, they react with _____ in the carbon dioxide buffer system.
12. Yawning is a mechanism by which the body gets rid of excess _____.
13. The pH of an acidic lake would be less than _____.
14. The two most common bases used in antacids are _____.
15. The long calibrated tube used in titrations is called a(n) _____.

The dissolved carbon dioxide blood buffer system maintains quite a constant pH in blood. Write the word or phrase that best completes each statement about this system.

16. The CO_2 dissolves in water to produce _____.
17. The other part of the buffer system is the _____ ion.
18. If OH^- ions increase in the blood, _____ reacts to lower its concentration.

A buffer system used by the body to control large changes in pH is the phosphate system. The system consists of the hydrogen phosphate (HPO_4^{2-}) and dihydrogen phosphate (H_2PO_4^-) ions of the triprotic (three-proton) acid H_3PO_4 .

19. Write an ionic equation that shows how the hydrogen phosphate ion reacts with a water molecule.

20. Write a net ionic equation that shows what happens when a hydrogen ion is added to the products of the reaction in question 19.

21. Write a net ionic equation that shows how a hydroxide ion reacts with a dihydrogen phosphate ion.

Answer each of the following questions about titration reactions.

22. In a titration, 50 mL of 0.2M sodium hydroxide reacts exactly with 250 mL of hydrochloric acid. What is the molarity of the hydrochloric acid?

23. In a titration between 0.1M hydrochloric acid and a base of unknown strength, 40 mL of the acid were required to neutralize exactly 8 mL of the base. What is the molarity of the base? _____

24. 23.8 mL of a sulfuric acid solution is required to neutralize 34.2 mL of a 0.2M potassium hydroxide solution. What is the molarity of the sulfuric acid solution? _____

25. How many milliliters of 2M sodium hydroxide are needed to neutralize 15 mL of 0.6M hydrochloric acid? _____

Name _____

Date _____

Class _____

16.1 The Nature of Oxidation-Reduction Reactions

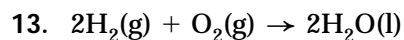
CHAPTER 16

Use with text pages 554 – 562

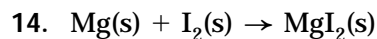
Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 1. Every oxidation reaction is accompanied by a reduction reaction.
- _____ 2. A compound formed between oxygen and one other element is called an acid.
- _____ 3. Oxidation is any reaction in which one element or compound loses electrons to a more electronegative element.
- _____ 4. One way to protect iron from rusting is to cover it with a less active metal, such as zinc.
- _____ 5. Another name for an oxidation-reduction reaction is a synthesis reaction.
- _____ 6. The oxidation number of any free element is 1+.
- _____ 7. When an element is oxidized, its oxidation number increases.
- _____ 8. Redox half-reactions seldom occur alone.
- _____ 9. A net oxidation-reduction reaction is the sum of the oxidation half-reaction plus the reduction half-reaction.
- _____ 10. During a redox reaction, an oxidizing agent is always reduced.
- _____ 11. Oxygen must be present for a redox reaction to occur.
- _____ 12. During a redox reaction, a reducing agent loses oxygen.

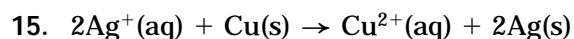
For each of the following reactions, on the space provided, identify the element that is oxidized (O), the element that is reduced (R), the oxidizing agent (OA), and the reducing agent (RA).



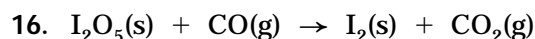
O _____ R _____ OA _____ RA _____



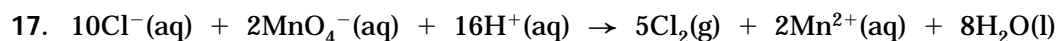
O _____ R _____ OA _____ RA _____



O _____ R _____ OA _____ RA _____

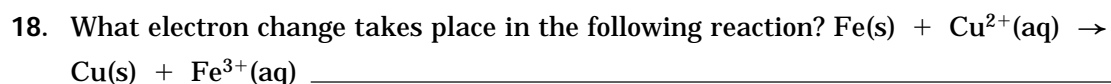


O _____ R _____ OA _____ RA _____



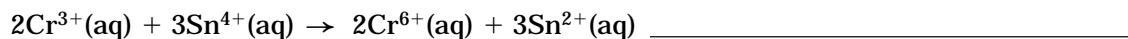
O _____ R _____ OA _____ RA _____

Answer the following questions.



19. What general rule can you state about the exchange of electrons between the oxidizing agent and the reducing agent in a redox reaction? _____

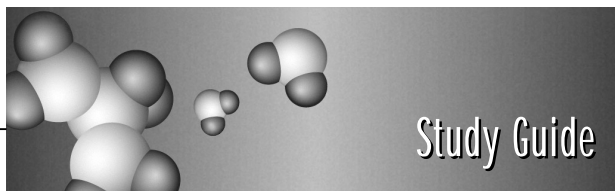
20. What is the total number of electrons exchanged in the following redox reaction?



21. Write the oxidation half-reaction and the reduction half-reaction for the following redox reaction. $2\text{Al}(\text{s}) + 6\text{HCl}(\text{aq}) \rightarrow 2\text{AlCl}_3(\text{aq}) + 3\text{H}_2(\text{g})$ _____

22. What is the oxidizing agent and the reducing agent in the following reaction?





17.1 Electrolysis: Chemistry from Electricity

CHAPTER 17

Use with text pages 584 – 598

Place a *T* for true or an *F* for false in front of each of the following statements. Then, for those statements that are false, change the underlined word or words to make the statement true.

- _____ 1. The first practical battery was constructed by Galvani.
- _____ 2. Electrolysis is a process by which an electrical current causes a spontaneous redox reaction to occur.
- _____ 3. During electrolysis, electrons flow into a cell through the cathode.
- _____ 4. In a Downs cell, sodium metal forms at the carbon anode.
- _____ 5. The ion HCO_3^- is an example of an anion.
- _____ 6. Charge is conducted through an electrolytic cell by means of the anode.
- _____ 7. The purpose of a battery in an electrolysis system is to pump electrons through the system.
- _____ 8. Every electrolytic cell must produce at least one half-reaction(s).
- _____ 9. The purpose of using hydroelectric power for commercial electrolysis operations is to ensure a sufficient supply of water for the operations.
- _____ 10. The substance from which aluminum is obtained in the Hall-Héroult process is aluminum oxide, Al_2O_3 .
- _____ 11. During the electroplating of an object, the object is used as the anode in an electrolytic cell.
- _____ 12. An electrochemical procedure for identifying large biological molecules is electrophoresis.

Answer the following questions.

13. Why do oxidation and reduction always occur together? _____

14. Is it possible to have an electrical current that consists of something other than electrons? _____

15. Since sodium chloride already exists in the form of ions, why is it necessary to melt the compound before decomposing it in a Downs cell? _____
- _____
- _____
16. When an aqueous solution of sodium chloride is electrolyzed, hydrogen gas is produced at the cathode. When an aqueous solution of copper(II) chloride is electrolyzed, copper is produced at the cathode. How do you explain this difference?
- _____
- _____
17. Why are the electrodes used in electrolytic processes such as in the Downs cell and in the Hall-Héroult process always made of inert substances, such as graphite or platinum. _____
- _____

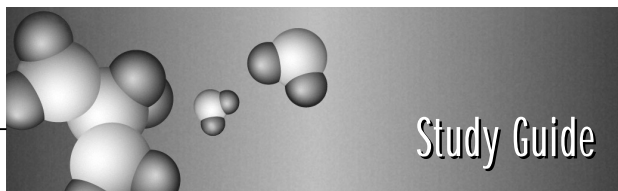
In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- _____ 18. The first person to observe the flow of electricity between two dissimilar metals was _____.
- a. Humphry Davy c. Alessandro Volta
- b. Luigi Galvani d. Charles H. E. Downs
- _____ 19. The part of an electrolytic cell at which reduction occurs is the _____.
- a. anode c. battery
- b. cathode d. electrolyte
- _____ 20. If 500 electrons per second are being released at one electrode of an electrolytic cell, then the number of electrons per second being used up at the other electrode is _____.
- a. at least 500 c. greater than 500
- b. exactly 500 d. dependent on the chemicals used
- _____ 21. The product that can be obtained by the electrolysis of molten sodium chloride but not of aqueous sodium chloride is _____.
- a. hydrogen gas c. metallic sodium
- b. chlorine gas d. oxygen gas

Name _____

Date _____

Class _____



Study Guide

17.2 Galvanic Cells: Electricity from Chemistry

CHAPTER 17

Use with text pages 599 – 615

Write the word or phrase that best completes each of the following sentences.

1. In a lemon battery, current is conducted through the lemon by means of _____.
2. The reaction that takes place at the anode of the lemon battery for any metal, M , is _____.
3. The term _____ is used to describe the difference in electrical “pressure” between any two electrodes in a galvanic cell.
4. Unlike galvanic cells, electrolytic cells have a(n) _____.
5. In a galvanic cell, electrical energy comes from _____.
6. Of the two metals, zinc and lead, _____ is more easily oxidized.
7. A half-reaction in which silver ions are reduced is _____.
8. If silver were used as an electrode in a galvanic cell in which the half-reaction in question 7 takes place, it would be the _____ of that cell.
9. In a galvanic cell consisting of a strip of copper, a strip of iron, electrolytes, and a salt bridge, the only way to increase the total amount of energy produced by the cell is _____.
10. One way to increase the potential difference in the galvanic cell in question 9 is to _____.
11. A dry cell is not really dry because the electrolyte it contains is in the form of a(n) _____.

12. In a dry cell, the zinc casing acts as the _____ of the cell.
13. The two electrodes in a lead storage battery are made of _____ and _____.
14. The product formed in a lead storage battery as the battery discharges is _____.

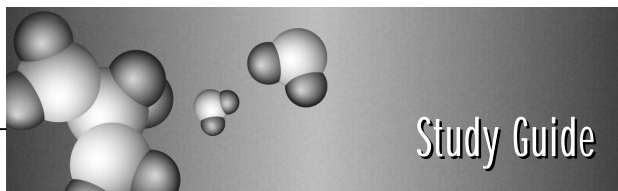
In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- _____ 15. If aluminum and iron are both present in a galvanic cell, the reaction most likely to take place is _____.
- a. $\text{Al} \rightarrow \text{Al}^{3+} + 3e^{-}$ c. $\text{Al}^{3+} + 3e^{-} \rightarrow \text{Al}$
b. $\text{Fe} \rightarrow \text{Fe}^{2+} + 2e^{-}$ d. $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + e^{-}$
- _____ 16. A galvanic cell that can be transported somewhat easily is known as a(n) _____.
- a. voltaic cell c. battery
b. electrolytic cell d. fuel cell
- _____ 17. The type of chemical reaction that always takes place within a galvanic cell is _____.
- a. synthesis c. double displacement
b. single displacement d. redox
- _____ 18. The amount of energy available from a galvanic cell with electrodes made of metals A and B depends on _____.
- a. the amount of A and B available c. neither of these factors
b. the potential difference between electrodes A and B d. both of these factors
- _____ 19. The fewest number of galvanic cells a battery can contain is _____.
- a. one c. six
b. two d. twelve

Name _____

Date _____

Class _____



18.1 Hydrocarbons

CHAPTER 18

Use with text pages 622 – 639

For each item in Column A, write the letter of the matching item in Column B.

Column A

- _____ 1. alkane
- _____ 2. alkene
- _____ 3. alkyne
- _____ 4. aromatic hydrocarbon
- _____ 5. cracking
- _____ 6. fractional distillation
- _____ 7. isomer
- _____ 8. reforming
- _____ 9. unsaturated hydrocarbon

Column B

- a. process that separates petroleum into many hydrocarbons
- b. process that converts large alkanes to aromatic hydrocarbons
- c. $\text{CH}_3\text{—CH=CH}_2$, $\text{CH}\equiv\text{CH}$, or $\text{CH}_3\text{—CH}_2\text{—CH=CH}_2$
- d. compounds with the same formula but different structures
- e. benzene
- f. decane
- g. $\text{CH}_3\text{—CH=CH—CH}_3$
- h. process that converts large alkanes to small alkanes and alkenes
- i. $\text{CH}_3\text{—CH}_2\text{—CH}_2\text{—C}\equiv\text{C—CH}_3$

Compare the hydrocarbons in the following pairs. Explain how they differ or state the relationship between the hydrocarbons.

- 10. alkanes and alkenes _____

- 11. alkenes and alkynes _____

- 12. straight and branched chains _____

- 13. aromatic and alkane ring compounds _____

- 14. petroleum and gasoline _____

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15. propane and hexane _____

16. cyclohexane and benzene _____

17. 1-butene and 2-butene _____

18. *trans*-2-butene and *cis*-2-butene _____

19. 2-methylpentane and 3-methylpentane _____

20. cyclopentane and cyclohexane _____

21. methane and methyl _____

22. saturated and unsaturated _____

23. 2,2-dimethylpentane and 2,3-dimethylpentane _____

24. benzene and naphthalene _____

Name _____

Date _____

Class _____

18.2 Substituted Hydrocarbons

CHAPTER 18

Use with text pages 640 – 647

For each functional group in Column A, write the letter of the correct organic family in Column B.

Column A

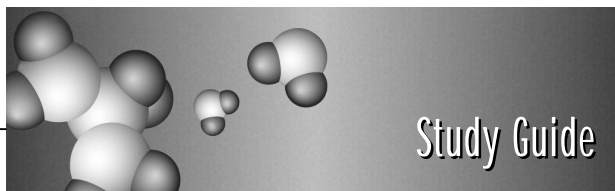
- _____ 1. $\begin{array}{c} \text{—C=C—} \\ | \quad | \end{array}$
- _____ 2. $\text{—C}\equiv\text{C—}$
- _____ 3. $\begin{array}{c} | \\ \text{—C—OH} \\ | \end{array}$
- _____ 4. $\begin{array}{c} | \\ \text{—C—Cl} \\ | \end{array}$
- _____ 5. $\begin{array}{c} \text{O} \\ || \\ \text{—C—OH} \end{array}$
- _____ 6. $\begin{array}{c} \text{O} \\ || \\ \text{R—C—O—R'} \end{array}$
- _____ 7. $\begin{array}{c} \text{O} \\ || \\ \text{R—C—R'} \end{array}$
- _____ 8. $\begin{array}{c} \text{O} \\ || \\ \text{R—C—H} \end{array}$
- _____ 9. R—O—R'
- _____ 10. R—NH_2
- _____ 11. $\begin{array}{c} \text{O} \\ || \\ \text{R—C—NH}_2 \end{array}$

Column B

- a. alkyne
- b. ether
- c. amide
- d. aldehyde
- e. alkene
- f. alcohol
- g. ketone
- h. amine
- i. carboxylic acid
- j. ester
- k. halogenated hydrocarbon

In the space at the left, write the letter of the class of compounds that best fits the property listed.

- _____ 12. Are found in proteins
- a. esters
- b. ethers
- c. aldehydes
- d. amides
- _____ 13. Have a tart taste, often with a strong, unpleasant odor
- a. aldehydes
- b. ketones
- c. carboxylic acids
- d. esters



18.3 Plastics and Other Polymers

CHAPTER 18

Use with text pages 648 – 661

Use each of the terms below just once to complete the passage.

addition reaction
condensation
cross-linking

monomer
polymer

thermoplastic
thermosetting

Polyethylene is a synthetic material used for many different purposes, including the manufacturing of milk jugs and plastic beakers. The repeating unit, called a (1) _____, which forms polyethylene, is ethylene. Many hundreds or thousands of ethylene molecules combine in a reaction known as a(n) (2) _____, in which double bonds are broken when the repeating units combine. Because it is made of so many repeating units, polyethylene is an example of a(n) (3) _____. In polyethylene, the molecules are arranged in a long, linear sequence. Such materials are usually (4) _____, meaning that they tend to melt or soften when heated. Other polymers do not melt when heated and are called, therefore, (5) _____ plastics. One reason that some polymers tend not to melt when heated is that their molecules undergo (6) _____, which tends to make the material more rigid. Other polymers are formed in a chemical reaction known as (7) _____ during which a small molecule such as water is eliminated during the reaction.

Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 8. Every polymer includes repeating units called monomers.
- _____ 9. Teflon is a naturally-occurring polymer of tetrafluoroethylene.
- _____ 10. The properties of a polymer are exactly like those of the monomer of which it is formed.
- _____ 11. DNA, protein, and silk are examples of naturally-occurring polymers.

- _____ 12. Cellulose and starch are made of the same monomer.
- _____ 13. Addition reactions cannot occur between double-bonded molecules.
- _____ 14. During vulcanization, sulfur is added to liquid latex to make a more flexible product.
- _____ 15. Polyethylene is an example of a thermoplastic polymer.

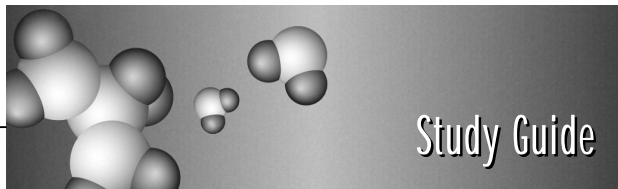
In each of the following statements, select the letter of the word or phrase that best completes the sentence.

- _____ 16. A polymer that was discovered accidentally is _____ .
- nylon
 - polyethylene
 - Teflon
 - rubber
- _____ 17. An example of a naturally occurring polymer is _____ .
- polyethylene
 - cellulose
 - Teflon
 - nylon
- _____ 18. The monomer from which cellulose and starch is made is _____ .
- glucose
 - ribose
 - cyclohexane
 - benzene
- _____ 19. Whether a monomer will undergo addition polymerization or condensation polymerization depends on _____ .
- the size of the molecule
 - the molecule's structure
 - whether the molecule is aromatic
 - whether the molecule occurs naturally or not
- _____ 20. When two polymer chains link to each other, the process is known as _____ .
- addition
 - condensation
 - cross-linking
 - chain-linking

Name _____

Date _____

Class _____



19.1 Molecules of Life

CHAPTER 19

Use with text pages 668 – 691

In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- _____ 1. The monomers from which proteins are made are _____.
 - a. simple sugars
 - b. amino acids
 - c. nitrogen bases
 - d. fatty acids

- _____ 2. _____ generally are proteins.
 - a. Cellulose
 - b. Amino acids
 - c. Lipids
 - d. Enzymes

- _____ 3. The linkage of atoms that joins the monomers in a protein molecule is known as a(n) _____.
 - a. peptide bond
 - b. active site
 - c. disulfide bond
 - d. hydrogen bond

- _____ 4. The point at which an enzyme bonds to a substrate is known as the _____.
 - a. peptide bond
 - b. active site
 - c. binding site
 - d. substrate groove

- _____ 5. Which of the following is a disaccharide?
 - a. glucose
 - b. fructose
 - c. ribose
 - d. sucrose

- _____ 6. The organic family to which glycerol belongs is the _____.
 - a. alcohols
 - b. alkanes
 - c. carboxylic acids
 - d. simple sugars

- _____ 7. A polyunsaturated fatty acid molecule contains _____.
 - a. one double bond
 - b. one triple bond
 - c. more than one double bond
 - d. both double and triple bonds

- _____ 8. One important way lipids are used in the body is _____.
 - a. to make proteins
 - b. as a source of energy
 - c. to digest fats
 - d. as enzymes

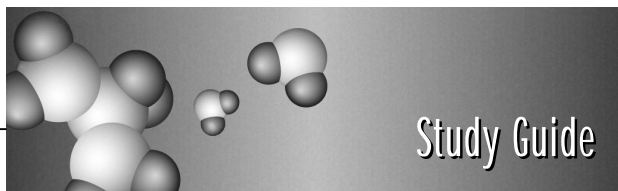
- _____ 9. A nucleotide consists of a simple sugar, a phosphate group, and _____.
 - a. a fatty acid
 - b. a disaccharide
 - c. glycerol
 - d. a nitrogen base

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Name _____

Date _____

Class _____



19.2 Reactions of Life

CHAPTER 19

Use with text pages 692 – 700

Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 1. The processes by which carbohydrates, proteins, and lipids are digested are catalyzed by disaccharides.
- _____ 2. The energy stored in the bonds of nutrient molecules was originally converted from light energy by means of respiration.
- _____ 3. Insulin is an example of a hormone, a molecule that carries messages throughout the body.
- _____ 4. Most of the energy used by cells comes from the oxidation of proteins.
- _____ 5. Chemical reactions that take place in cells occur as rapidly as they do because of the presence of lipids in the cells.
- _____ 6. From a chemical standpoint, respiration is the process by which glucose reacts with oxygen to form water and carbon dioxide.
- _____ 7. When a phosphate group is added to ADP, energy is released.
- _____ 8. In the final step of the electron transport chain, electrons are transferred to oxygen atoms.
- _____ 9. Cells die when deprived of oxygen because they can no longer harness electrons.
- _____ 10. The fatigue in an athlete's muscles after exercising is caused by the presence of ethanol.
- _____ 11. During glycolysis, 2 molecules of ATP are produced.
- _____ 12. The total number of ATP molecules produced during the breakdown of glucose is 36.

Provide the missing part of each of the following chemical equations, and tell what function the reaction has in the body.

13. glucose + oxygen → _____

14. _____ → ATP _____

15. $\text{NAD}^+ + e^- \rightarrow$ _____

16. three-carbon compound + yeast → _____ + carbon dioxide

Answer the following questions.

17. Name the types of organic compounds you would expect to find in the bloodstream.

18. Explain whether it is correct to say that the term *metabolism* refers to the chemical reactions by which biomolecules are broken down to produce energy.

19. What is the connection between breathing and respiration? _____

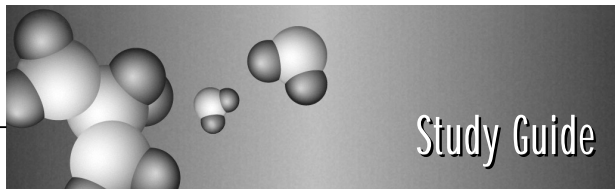
20. Where does the oxygen come from to make the water that is the final product of respiration? Where does the hydrogen come from? _____

21. What is the function of alcoholic fermentation in the baking industry?

Name _____

Date _____

Class _____



20.1 Energy Changes in Chemical Reactions

CHAPTER 20

Use with text pages 708 – 718

Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- Heat is released in an exothermic reaction.

- Combustion is a familiar example of a(n) endothermic reaction.

- The unit used to measure energy in the SI system is the Calorie.

- According to the law of conservation of energy, energy cannot be changed from one form to another during a chemical reaction.

- The value of ΔH for exothermic reactions is always negative.

- Energy is conserved only in exothermic reactions.

- Activation energy is generally required to initiate endothermic, but not exothermic reactions.

- In an exothermic reaction, the energy of the products is less than the energy of the reactants.

- The use of a catalyst increases the amount of product formed.

- An increase in entropy corresponds to a(n) increase in disorder.

- Exothermic reactions are usually spontaneous.

- The total amount of entropy in the universe has a tendency to stay the same.

13. In a spontaneous reaction, entropy can decrease.

14. An endothermic reaction requires a continuous input of energy.

Answer the following questions.

15. Chemical systems move toward decreased energy. Does that mean that all chemical reactions are exothermic? Explain. _____

16. Why does steam have greater entropy than liquid water? _____

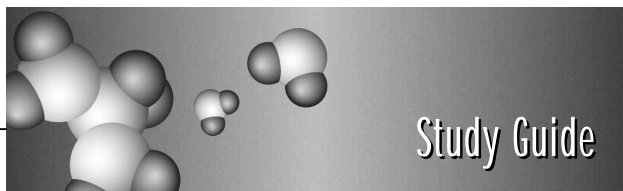
17. Which has greater entropy, a tablespoon of dry salt or a tablespoon of salt dissolved in water? Explain. _____

18. Assuming that temperature does not change, would there be an increase or decrease in entropy in the following reaction $\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_3(\text{aq})$? Explain. _____

Name _____

Date _____

Class _____



20.2 Measuring Energy Changes

CHAPTER 20

Use with text pages 719 – 737

For each term or symbol in Column A, write the letter of the matching definition in Column B.

Column A

- _____ 1. ΔT
- _____ 2. C_w
- _____ 3. q_w
- _____ 4. q_{reaction}
- _____ 5. 1 kJ
- _____ 6. 1 cal
- _____ 7. 1 J
- _____ 8. Cal
- _____ 9. $^{\circ}\text{C}$
- _____ 10. g/mol

Column B

- a. heat gained or lost by water
- b. 1000 cal
- c. 4.184 J
- d. molar mass
- e. temperature change
- f. Celsius temperature
- g. heat given off or absorbed in a chemical change
- h. 0.239 cal
- i. specific heat of water
- j. 1000 joules

Solve the following problems.

11. A sample of butane (C_4H_{10}) gas is burned inside a calorimeter that contains 2.0 kg of water. The temperature of the water rises by 10.2°C . How much heat was produced in the reaction? _____
12. A 5.5-g sample of steel wool is ignited in a calorimeter containing 1.75 kg of water. The temperature of the water rises from 22.50°C to 29.30°C . What is the heat produced by this chemical change? _____
13. Convert each of the following measurements.
 - a. 2.38 cal to J _____
 - b. 3560 cal to kcal _____
 - c. 165.9 kJ to Cal _____
 - d. 8191 kJ to cal _____

14. The data below was obtained for a sample of peas burned in a calorimeter. Find the energy released, in both kilojoules per gram and Calories per gram.

Mass 2.75 grams

Mass of water in calorimeter 1500 grams

Initial temperature ($^{\circ}\text{C}$) 20.00 $^{\circ}\text{C}$

Final temperature ($^{\circ}\text{C}$) 23.00 $^{\circ}\text{C}$

Energy content (kJ/g; Cal/g) _____

Answer the following questions.

15. What is a calorimeter? Explain how the device is used. _____

16. Explain the relationship between calories, Calories, and kilocalories.

17. What is efficiency? Explain why it is impossible for any energy-related process to be 100 percent efficient. _____

In the space provided, explain how the two terms in the following pairs differ from each other or are related to each other.

9. photosynthesis and respiration _____

10. chlorophyll and chloroplasts _____

11. glucose and carbohydrate _____

12. light reactions and Calvin cycle _____

Answer the following questions.

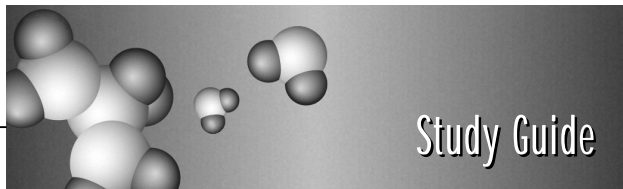
13. A classical question in the study of photosynthesis was whether the oxygen released during the process comes from carbon dioxide or water. What is now known to be the correct answer to that question? _____

14. Microorganisms known as purple bacteria capture solar energy by means of a process similar to photosynthesis, but they use hydrogen sulfide (H_2S) instead of water. What would be the products of this process? _____

Name _____

Date _____

Class _____



21.1 Types of Radioactivity

CHAPTER 21

Use with text pages 744 – 760

For each item in Column A, write the letter of the matching item in Column B.

Column A

- _____ 1. ${}^4_2\text{He}$
- _____ 2. ${}^0_{-1}e$
- _____ 3. γ
- _____ 4. Becquerel
- _____ 5. Curie
- _____ 6. alpha decay
- _____ 7. beta decay
- _____ 8. phosphorescence
- _____ 9. ${}^9_4\text{Be}$
- _____ 10. ${}^{234}_{90}\text{Th}$
- _____ 11. scintillation
- _____ 12. ${}^{14}_6\text{C}$

Column B

- a. gamma ray
- b. coined the term *radioactivity*
- c. process in which atomic number decreases by 2
- d. ${}^{238}_{92}\text{U} \rightarrow {}^4_2\text{He} + ?$
- e. ability to continue glowing after absorbing light
- f. a flash of light
- g. is called an alpha particle
- h. is used to determine age of once-living materials
- i. discoverer of radioactivity
- j. contains four protons and five neutrons
- k. process in which atomic number increases by 1
- l. is called a beta particle

Answer the following questions.

13. Contrast alpha decay, beta decay, and gamma decay in terms of what is released and the change in the atomic number and mass number of the nucleus.

14. Explain what a Geiger counter is used for and how it operates. _____

15. Explain what is meant by the half-life of a radioactive isotope. _____

Solve the following problems.

16. After 8.0 years, how much remains of a radioactive isotope whose half-life is 2.0 years if the original mass of the sample was 32 g? _____

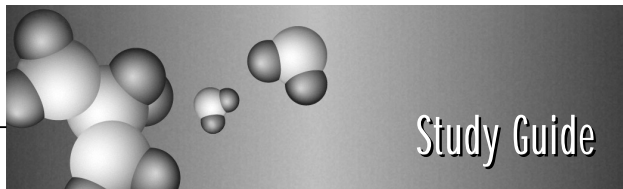
17. How long will it take a 60.0 g-sample of a radioactive isotope with a half-life of 4.30×10^6 years to decay to 7.5 g? _____

18. A sample of a radioactive isotope with a mass of 64.0 g decays to 1.0 g in 18 minutes. What is the half-life of the isotope? _____

19. It has taken a radioactive isotope with a half-life of 3.24 years 16.2 years to reduce in mass to 2.5 g. What was the original mass of the isotope?

Name _____

Date _____ Class _____



21.2 Nuclear Reactions and Energy

CHAPTER 21

Use with text pages 761 – 767

Place a *T* for true or an *F* for false on the blank for each of the following statements. Change the underlined word or phrase in each false statement to make it true. Write your correction on the blank.

- _____ 1. The equivalence of mass and energy was discovered by Albert Einstein.
- _____ 2. The first person to observe a nuclear fission reaction was Lise Meitner.
- _____ 3. The subatomic particle that is necessary in order to initiate the nuclear fission of uranium-235 is a(n) alpha particle.
- _____ 4. The subatomic particle that is produced during a nuclear chain reaction involving uranium-235 is a(n) neutron.
- _____ 5. The most common isotope used in nuclear reactors today is uranium-238.
- _____ 6. The purpose of the control rods in a nuclear reactor is to release neutrons produced during nuclear fission.
- _____ 7. Steam produced in a nuclear reactor is used to drive a(n) steam reactor in a nuclear power plant.
- _____ 8. Currently, the only element other than uranium that is used as a fuel in nuclear reactors is plutonium.
- _____ 9. A deuterium atom differs from an atom of hydrogen-1 in the number of protons in its nucleus.
- _____ 10. One type of fusion reactor currently being tested is called the tomahawk.

State the function of each of the following parts of a nuclear power plant.

11. Fuel rods _____

12. Control rods _____

13. Moderator _____

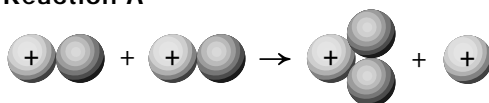
14. Coolant _____

15. Turbine _____

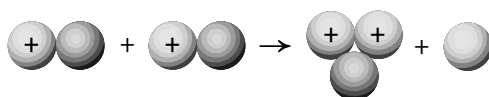
16. Generator _____

The illustration below shows two nuclear reactions that involve the same reactant. Answer the following questions about these reactions.

Reaction A



Reaction B

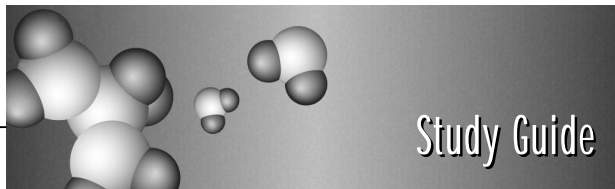


17. What kind of nuclear reaction is illustrated in each reaction? _____

18. Write the symbol for the reactant, using the notation that shows the mass number and atomic number. What is the name of this isotope? _____

19. Write the symbols for the two products in Reaction A. what are the names of these particles? _____

20. Write the symbols for the two products in Reaction B. What are the names of these products? _____



21.3 Nuclear Tools

CHAPTER 21

Use with text pages 768 – 779

In the space at the left, write the letter of the word or phrase that best completes the statement or answers the question.

- _____ 1. When radiation is used to treat a disease, the radiation destroys _____ .
- a. only cancer cells c. both cancer and healthy cells
b. only healthy cells d. only rapidly growing cells
- _____ 2. When radioactive isotopes are used for diagnostic purposes, they are being used to _____ .
- a. treat a disease c. learn more about a normal function
b. identify a medical disorder d. kill cancer cells
- _____ 3. The correct nuclear symbol for a positron is _____ .
- a. ${}^+_0 e$ c. ${}^0_{-1} e$
b. ${}^{-1}_0 e$ d. ${}^0_{+1} e$
- _____ 4. The letter *m* in the symbol for the radioactive isotope technetium-99m indicates that when a nucleus of this isotope decays, its atomic mass will _____ .
- a. decrease by four c. decrease by two
b. increase by two d. stay essentially the same
- _____ 5. Cancer cells are more likely to be killed by radiation than are healthy cells because the cancer cells _____ .
- a. are more numerous c. divide more rapidly
b. contain DNA d. divide more slowly
- _____ 6. The only kind of isotope listed below that could not be used as a tracer is _____ .
- a. an alpha emitter c. a beta emitter
b. a gamma emitter d. a stable isotope
- _____ 7. Radiation can be used to preserve foods because it destroys _____ .
- a. microorganisms c. DNA in food cells
b. toxic chemicals d. decayed food cells
- _____ 8. A device in which artificial elements can be made is a(n) _____ .
- a. PET machine c. cyclotron
b. Geiger counter d. gamma irradiator

