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Chapter Outline

1.1 WHAT IS CHEMISTRY?
1.2 THE SCIENTIFIC METHOD
1.3 INTRODUCTION TO CHEMISTRY
1.1 What is Chemistry?

Lesson Quiz

Name___________________ Class________________ Date________

Multiple Choice:

1. Copernicus studied
   a. anatomy
   b. painting
   c. movement of the sun
   d. movement of the planets

2. Our major source of energy today is
   a. wind
   b. fossil fuels
   c. nuclear energy
   d. solar power

3. Genetic information is contained in
   a. proteins
   b. enzymes
   c. DNA
   d. lipids

4. Which one of the following is not a material created by chemists?
   a. liquid crystals
   b. enzymes
   c. plastics
   d. ceramics

5. Chemists contribute to the growth of food by all of the following except
   a. developing new fertilizers
   b. causing soil erosion
   c. creating new pesticides
   d. making new soil supplements

6. Algae blooms in water are caused by
   a. excess oxygen in the water
   b. death of fish
   c. decrease in oxygen in the water
   d. fertilizer run-off from the ground

7. Many scientists believe that global warming is caused by
   a. sun spots
   b. increase in carbon dioxide from burning fossil fuels
c. lowered carbon dioxide due to increased use by plants
d. forest fires

8. Improved rubber compounds for tires might be developed by
   a. an inorganic chemist
   b. a biochemist
   c. an organic chemist
   d. a physical chemist

9. There is often overlap between the disciplines of biochemistry and
   a. inorganic chemistry
   b. analytical chemistry
   c. organic chemistry
   d. physical chemistry

10. All of the following are body parts developed by chemists except
    a. hip replacement
    b. artificial skin
    c. plastic blood vessels
    d. artificial liver

True/False:

11. _____ Microscopic refers to the small particles that make up all matter.
12. _____ Observing the rusting of iron is a microscopic process.
13. _____ Chemistry is the study of planetary orbits.
14. _____ The alchemists were never successful in their attempts to make gold.
15. _____ The elixir of life was supposed to convey immortality to humans.
16. _____ Pure chemistry always has a practical goal.
17. _____ The development of wrinkle-free fabrics is an example of applied chemistry.
18. _____ Studying how fast crystals form from salt solutions would be done by a physical chemist.
19. _____ A biochemist would do research on glucose use by the liver.
20. _____ Mercury is not toxic to humans.

Fill in the blanks:

21. Chemistry is the study of the ________________ of matter and the ____________ that matter undergoes.
22. Matter is anything that has ___________ and takes up space.
23. A ________________ would study the structure of the hemoglobin molecule and how it transports oxygen.
24. An organic chemistry works mainly with ____________ compounds.
25. The ____________ chemist would be interested in the analysis of rubies.
26. ____________ chemistry is the study of the composition of matter, with a focus on separating, identifying,
   and quantifying chemical samples.
27. One element known to cause brain damage in children is ____________.
28. Nuclear ________ is a process that occurs in the sun and stars.
29. Alchemy contributed to the production of ____________.
30. The first attempts to classify substances were made by ____________.
31. Iron is caused to rust by exposure to _________ and ____________.

Short Answers:

32. Rock salt sprinkled on ice will cause the ice to melt. Is this a macroscopic or microscopic process?
33. Some scientists study the chemical processes that take place in the sun. Is this pure or applied chemistry? Explain your answer.

34. Why are some people concerned about using nuclear power plants to generate electricity?

Answer Key

1. d
2. b
3. c
4. b
5. b
6. d
7. b
8. c
9. c
10. d
11. true
12. false
13. false
14. true
15. true
16. false
17. true
18. true
19. true
20. false
21. composition, changes
22. mass
23. biochemist
24. carbon
25. inorganic
26. analytical
27. lead
28. fusion
29. gunpowder
30. alchemists
31. rainwater, oxygen in air
32. macroscopic - we can observe the melting directly
33. pure chemistry since there is no immediate application. However, some aspects could be applied to develop new energy sources.
34. Nuclear power can generate a lot of electricity. The drawback is that the nuclear wastes remain radioactive for a long period of time and are hazardous.
Lesson Quiz

Name___________________ Class______________ Date________

Multiple Choice:

1. Copernicus made the following contribution to scientific progress
   a. research on the behavior of solar flares
   b. measured the path of the moon
   c. developed the idea of the heliocentric solar system
   d. discovered the double helix

2. The Renaissance was
   a. a time of scientific progress
   b. a medieval fair
   c. a time when science was not important
   d. a form of music

3. Experiments are important because they
   a. provide jobs for scientists
   b. test the laws of science
   c. test the hypothesis
   d. identify the control group

4. A theory has been
   a. repeatedly tested and shown to be true
   b. repeatedly tested without clear conclusions
   c. tested once and accepted
   d. poorly accepted by other scientists

5. All but one statement below tell why scientists work in groups
   a. problems are complex
   b. research progress takes a lot of time and money
   c. some scientists are lazy and steal the ideas of others
   d. research progress occurs in small steps, so many people are needed

6. What does peer review do for science?
   a. allows reviewers to eliminate research they don’t agree with
   b. makes sure research papers have reliable information
   c. costs extra money
   d. helps get a textbook written

7. A scientific law
   a. is always true
   b. might be disproved at some point in the future
1.2. The Scientific Method

- c. is another term for a hypothesis
- d. does not have a lot of data to support it

8. The independent variable
   - a. always increases with an increase in temperature
   - b. is never affected by temperature
   - c. is changed during the experiment
   - d. stays the same throughout the experiment

9. In order to accept the results of an experiment
   - a. other scientists must have failed to repeat it
   - b. it must be repeatable
   - c. a report needs to be written
   - d. only a few people need to agree with the data

10. The control group for an experiment
    - a. checks the reproducibility of the experiment
    - b. serves as the dependent variable
    - c. serves as the independent variable
    - d. does not experience the variable being tested

True/False:

11. _____ A theory has been widely tested
12. _____ A theory never changes
13. _____ The control group also has the experiment run on it
14. _____ The Renaissance put an emphasis on observation and experimentation
15. _____ The National Science Foundation is the only agency that funds research.
16. _____ The dependent variable is the one that is changed during the experiment.
17. _____ If experimental results are inconsistent with a hypothesis, the hypothesis must be changed or discarded.
18. _____ An experiment must only be repeated by the group originally proposing the hypothesis in order for the hypothesis to be accepted.
19. _____ Publishing research in a scientific journal is not the only way to get the information out to others.
20. _____ A valid explanation for an observation does not need to be tested if it sounds reasonable.

Matching:

21. _____ I develop an initial explanation that answers the question I ask.
22. _____ My final explanation has been repeatedly tested and is accepted by many scientists.
23. _____ I see something that is puzzling and that I cannot explain.
24. _____ what I get when I run an experiment
25. _____ variable observed during an experiment
26. _____ no known exceptions

a. observation
b. theory
c. hypothesis
d. dependent
e. results
f. law

Fill in the blanks:
27. The Renaissance was a time of great _________ in Europe.
28. Leonardo da Vinci carried out systematic studies of ________ and ________.
29. Great advancements in ____, ____ , and ____ were made during the Renaissance period.
30. The ________ model said that the Earth revolved around the Sun.
31. The scientific method is a _________ and __________ approach to the acquisition of knowledge.

**Critical Write Question:**

32. I want to study the effect of chocolate chips on the baking time of cookies. My experiment involves mixing different amounts of chocolate chips into cookie dough and baking the cookies in the oven at 350 °F for 30 minutes. I will assess the effect by observing the hardness of the final product.

   a. What is the independent variable in this experiment? Explain your answer.
   b. How many dependent variables are there? List them and explain your answer.
   c. Should I carry out my cookie research by myself or have other people help? Explain your answer.
   d. Briefly describe two ways I can share my cookie research with other scientists.

**Answer Key**

1. c
2. a
3. c
4. a
5. c
6. b
7. a
8. c
9. b
10. d
11. true
12. false
13. false
14. true
15. false
16. false
17. true
18. false
19. true
20. false
21. c
22. b
23. a
24. e
25. d
26. f
27. upheaval.
28. movement and aerodynamics.
29. art, music, culture.
30. heliocentric
31. systematic, logical
a. The independent variable is the amount of chocolate chips because we choose to vary this amount in each experiment.

b. There are three dependent variables: the cookie dough recipe, the temperature of the oven, and the length of time cookies are baked. Each of these is held constant throughout the experiment.

c. It might be better to have others help with these experiments. They can be done much more quickly and you have other people helping decide how hard the cookies are.

d. The cookie results can be presented at a scientific meeting or they can be written up in an article that would be published in a scientific journal.
1. Baking soda produces _____ during cooking.
   a. air
   b. carbon dioxide
   c. water
   d. hydrogen gas

2. Which of the following is not a part of the macroscopic world?
   a. rainwater
   b. rusting iron
   c. bacteria
   d. leaves

3. What was the philosopher’s stone?
   a. a material that protected the body
   b. a rock that Aristotle sat on while he taught
   c. a material used to refine ores
   d. a substance that could change lead into gold

4. Which of the following would not be considered applied chemistry?
   a. study of the atomic structure of rubidium
   b. use of rubidium in jewelry
   c. treatment of disease using rubidium solutions
   d. research on how to dissolve rubidium

5. Which of the following statements about fossil fuels is incorrect?
   a. coal, petroleum, and natural gas are fossil fuels
   b. they are a nonrenewable energy source
   c. the global supply of fossil fuels will never be used up
   d. fossil fuels supply most of our energy needs today

6. Which of the following statements is correct?
   a. plastics are never used in hip replacement because they wear out quickly
   b. genetically modified corn requires more pesticides than the non-modified plant
   c. plastic tubing has been successfully used to repair diseased blood vessels
   d. drugs do not affect any chemical processes in the body

7. Which one of the following is a material that was not developed by chemists?
   a. quartz
   b. adhesives
1.3. Introduction to Chemistry

8. Nuclear fusion
   a. is used to generate electrical energy
   b. takes place in the sun and stars
   c. involves making large molecules from smaller ones
   d. takes place at low temperatures

9. The dependent variable
   a. does not change as the temperature increases
   b. will change as the independent variable changes
   c. never changes
   d. is varied in the experiment to see how it influences the results

10. One of the following statements about Leonardo da Vinci is not true
    a. he was called the father of modern science
    b. he studied water flow
    c. he discovered penicillin
    d. he did medical dissections

True/False:

11. _____ A tablet of ibuprofen is macroscopic.
12. _____ Applied chemistry is involved in looking for basic scientific ideas.
13. _____ The alchemists wanted to find the philosopher’s stone.
14. _____ A biochemist studies how the pancreas affects blood glucose levels.
15. _____ A physical chemist looks at the composition of rocks.
16. _____ Nuclear energy is safer than using fossil fuels.
17. _____ Mercury has been shown to be a toxic element.
18. _____ Algae in water is a result of too little fertilizer.
19. _____ Peer review helps make sure scientific papers are reliable.
20. _____ Grant money can be used to purchase test tubes.

Fill in the blanks:

21. _____ is a toxic element other than lead that is no longer used to any extent.
22. Drugs work because of their _________ with other chemicals in the body.
23. Kevlar was discovered by __________ _______________ .
24. Alchemists laid the foundation for the production of _________________ .
25. Organic chemistry and _________ are often closely connected.
26. List the five points of the scientific method in order:
   a. _________
   b. _________
   c. _________
   d. _________
   e. _________
27. When scientists write grant proposals, they explain the _____ and _______ ______ of their research.

Short Answers:

28. Why is peer review an important part of the publication of scientific results?
29. Give an example in the medical field where a scientist might use more than one chemistry discipline.
Answer Key

1. b  
2. c  
3. d  
4. a  
5. c  
6. c  
7. a  
8. b  
9. b  
10. c  
11. true  
12. false  
13. true  
14. true  
15. false  
16. false  
17. true  
18. false  
19. true  
20. true  
21. mercury  
22. interactions  
23. Stephanie Kwolek.  
24. gunpowder.  
25. biochemistry  
   a. formulate a question  
   b. make a hypothesis  
   c. test the hypothesis through experimentation  
   d. analyze results  
   e. develop theories or revise hypothesis  
26. goals, projected costs  
27. to assure that the research is correct.  
28. A biochemist might identify a defect in the behavior of an organ. He/she would use organic chemistry to produce a drug for the medical problem and then use analytical chemistry to measure changes in the blood or other materials to study how effective the drug is.
CHAPTER 2

Matter and Change Assessments

Chapter Outline

2.1 PROPERTIES OF MATTER
2.2 CLASSIFICATION OF MATTER
2.3 CHANGES IN MATTER
2.4 MATTER AND CHANGE
Lesson Quiz

Name___________________ Class________________ Date________

Multiple Choice:

1. One of the following is not a physical property of iron.
   a. malleability
   b. magnetism
   c. rusting
   d. color

2. You have a sample of cobalt chloride with the following properties. Which one is an intensive property?
   a. 654 grams
   b. density of 3.356 g/cm³
   c. melting point of 735°C
   d. sky blue color

3. The ability of aluminum to be formed into sheets is called
   a. malleability
   b. ductility
   c. conductivity
   d. deformability

4. Which of the following is not a characteristic of a sample of sodium metal?
   a. defined mass
   b. defined volume
   c. takes on shape of container
   d. defined color

5. An irreversible change would be
   a. chipping ice to flakes
   b. getting a hair cut
   c. condensing steam
   d. boiling water

6. The particles of a solid have all the following properties except one
   a. very close together
   b. in fixed position
   c. easily compressed
   d. may not expand when heated

7. Which of the following is not a physical change that a lead bar could undergo?
   a. melting
   b. boiling
2.1. Properties of Matter

8. A gas sample will not demonstrate one of the following properties
   a. has a definite volume
   b. has no definite shape
   c. has a definite mass
   d. be easily compressed

9. A pure substance has the following characteristic
   a. uniform composition
   b. composition not definite
   c. can be compressed
   d. molecules rigidly arranged

10. One of the following statements is not true of a physical change
    a. the shape is unchanged
    b. the chemical composition is unchanged
    c. the identity of the material is unchanged
    d. the reactivity is unchanged

True/False:

11. _____ Bromine becomes another substance when it changes from liquid to gas.
12. _____ Magnetism is a physical property.
13. _____ A sample of liquid bromine cannot evaporate.
14. _____ Metallic sodium can become a vapor.
15. _____ The volume of a sample of water does not change when heated.
16. _____ Physical changes are always reversible.
17. _____ Air can be considered a pure substance.
18. _____ Physical changes involving a change of state are reversible.
19. _____ Burning of gasoline is not a physical change.
20. _____ Chopping down a tree is a reversible physical change.

Fill in the Blank:

21. Making furniture from wood is a(n) ________________ physical change.
22. A ____________ takes the shape of its container.
23. A ____________ forms when mercury is boiled.
24. The freezing point of bromine is a ____________ property.
25. The rich red color of a ruby is an ____________ property while the mass of the ruby is an ____________-
    property.
26. __________ is the amount of space occupied by a material.
27. A ____________ has a definite volume and shape.
28. _____ is the term used for a substance which has neither a definite shape nor a definite volume at room
    temperature.
29. The ability of copper to be made into sheets is called ________________.
30. __________, __________, and ____________ are excellent conductors of electricity.
31. How does the shape of a water sample change when it goes from liquid to solid?
32. Iodine has a melting point of 113.7°C and a boiling point of 184.3°C. Give the state an iodine sample is in at
    the following temperatures:
    a. 205.3°C __________
Answer Key

1. c
2. d
3. a
4. c
5. b
6. c
7. d
8. a
9. a
10. a
11. false
12. true
13. false
14. true
15. false
16. false
17. false
18. true
19. true
20. false
21. irreversible
22. liquid
23. vapor
24. physical
25. intensive, extensive
26. volume
27. solid
28. gas
29. malleability
30. silver, gold, copper
31. Water changes from a liquid that takes the shape of its container to a solid that has a definite shape.
   a. gas
   b. solid
   c. liquid
2.2 Classification Of Matter

Lesson Quiz

Name___________________ Class_________________ Date________

Multiple Choice:

1. A chemical symbol is
   a. something students need to memorize
   b. a unique one- or two-letter abbreviation for an element
   c. something that always refers to a Greek or Latin name
   d. a letter than can refer to several elements

2. The process of breaking a compound down into its elements is called
   a. recombination
   b. rearrangement
   c. dessication
   d. decomposition

3. Which of the following is an element?
   a. sea water
   b. salt
   c. a gold ring
   d. methane gas

4. Sodium chloride is changed to metallic sodium and chlorine gas. This process represents a
   a. physical change
   b. chemical change
   c. nuclear change
   d. elemental change

5. Chicken noodle soup is a __________ mixture.
   a. homogeneous
   b. heterogeneous
   c. homogenized
   d. heterophilic

6. Distillation is a process of separating materials by
   a. differences in solubility
   b. differences in particle size
   c. differences in color
   d. differences in boiling point

7. \( \text{H}_2\text{SO}_4 \) is a(n)
   a. element
   b. mixture
c. compound
d. heterogeneous solution

8. A chemical formula
   a. shows the elements in a compound
   b. shows how to make a compound
   c. shows the mixture present
   d. shows the decomposition’s products

9. Filtration separates
   a. two or more liquid phases
   b. solids in a heterogeneous mixture
   c. solids from liquids in a heterogeneous mixture
   d. solids from liquids in a homogeneous mixture

10. Another term for a homogeneous mixture is a
    a. solution
    b. suspension
    c. solubilizer
    d. stabilizer

True/False:

11. _____ One element can be converted to another element using chemical processes.
12. _____ A mixture can only have two components.
13. _____ Potassium iodide (KI) is a mixture of potassium and iodine.
14. _____ A chemical symbol for an element must have only one letter.
15. _____ Blood is a heterogeneous mixture.
16. _____ The compound NiS is composed of nickel and sodium.
17. _____ Cooking oil is a homogeneous mixture.
18. _____ A homogeneous mixture consists of one phase.
19. _____ Elements can be broken down into simpler substances.
20. _____ The two or more elements in a compound have been chemically combined.

Fill in the blank:

21. Sucrose (C\textsubscript{12}H\textsubscript{22}O\textsubscript{11}) is composed of ____________, ________________, and ____________.
22. The chemical symbol for silver is _____ and the Latin name is ________________.
23. Cu/Sn would represent a mixture of _____ and ____________.
24. Sushi is a ____________ mixture of rice and fish.
25. Kalium is the Latin name for the element ____.

Matching:

26. _____ change that produces matter with a different composition
27. _____ a mixture that is not uniform throughout.
28. _____ separate solid from liquids in a heterogeneous mixture.
29. _____ has uniform composition and properties

a. heterogeneous
b. phase
c. chemical change
d. filtration
30. Listed below are the boiling points for several petroleum fractions. Describe how you would separate diesel oil from everything else.

**Table 2.1:**

<table>
<thead>
<tr>
<th>Fraction</th>
<th>Boiling Point (°C)</th>
<th>Use</th>
</tr>
</thead>
<tbody>
<tr>
<td>butane</td>
<td>below 30</td>
<td>natural gas</td>
</tr>
<tr>
<td>gasoline</td>
<td>100-150</td>
<td>car and truck fuel</td>
</tr>
<tr>
<td>diesel oil</td>
<td>200-300</td>
<td>fuel for trucks and trains</td>
</tr>
<tr>
<td>greases and waxes</td>
<td>400-500</td>
<td>lubricants</td>
</tr>
</tbody>
</table>

31. Describe how you would separate a mixture of lead buckshot and potassium nitrate fertilizer.

**Answer Key**

1. b  
2. d  
3. c  
4. b  
5. b  
6. d  
7. c  
8. a  
9. c  
10. a  
11. true  
12. false  
13. false  
14. false  
15. true  
16. false  
17. true  
18. true  
19. false  
20. true  
21. carbon, hydrogen, oxygen  
22. Ag, argentum  
23. copper, tin  
24. heterogeneous  
25. potassium  
26. c  
27. a  
28. d  
29. b  
30. Heat the mixture to approximately 175°C to remove the butane and gasoline. Then heat the solution to approximately 350°C and condense the vapor to obtain the diesel oil.  
31. Add water to the mixture, dissolving the potassium nitrate. Filter the mixture to remove the lead buckshot, then evaporate the water to obtain the potassium nitrate.
Lesson Quiz

Name___________________ Class______________ Date________

Multiple Choice:

1. Which of the following is a physical change?
   a. oxidation of sulfur
   b. melting of sulfur
   c. reduction of sulfur
   d. combustion of sulfur

2. One chemical property of the element sodium is
   a. a density of 0.968 grams/cm$^3$
   b. a melting point of 97.8°C
   c. forms hydrogen gas when added to water
   d. becomes transparent at extremely high pressures

3. Which of the following energy changes is associated with a chemical reaction?
   a. condensation of steam
   b. melting of lead
   c. compressing of air
   d. heating of mercuric oxide

4. One of the following is not a sign of a chemical reaction
   a. change of color
   b. formation of a precipitate
   c. release of energy
   d. production of a gas

5. Lead iodide produces a _______ color.
   a. yellow
   b. orange
   c. red
   d. tan

6. _____ is an example of a chemical change
   a. melting
   b. rusting
   c. boiling
   d. freezing

7. Terms that are not used to describe chemical changes include
   a. fermenting
   b. condensing
2.3. Changes in Matter

- c. burning
- d. rotting

8. A more general term for the rusting process is
   - a. erosion
   - b. explosion
   - c. corrosion
   - d. collusion

9. Zinc reacts with hydrochloric acid to form
   - a. oxygen
   - b. hydrogen
   - c. chlorine
   - d. helium

10. The reaction of magnesium with oxygen is called
    - a. reduction
    - b. fermentation
    - c. decomposition
    - d. combustion

True/False:

11. _____ The reactants are listed on the right hand side of the arrow in a chemical equation.
12. _____ Minor products can be omitted from a chemical equation.
13. _____ The conversion of mercuric oxide to mercury produces energy during the reaction.
14. _____ Iron reacts with oxygen to form rust.
15. _____ Nitrogen gas forms when lead nitrate and potassium iodide react.
16. _____ In the reaction between zinc and sulfur, zinc is a reactant.
17. _____ Chemical properties can only be observed after the reaction has occurred.
18. _____ If zinc and sulfur are mixed at room temperature, no reaction occurs.
19. _____ No color change takes place when zinc and sulfur react.
20. _____ Light is emitted when mercuric oxide decomposes.

Fill in the Blank:

21. In the reaction involving the electrolysis of water, we can write the following equation:
   
   \[ 2H_2O \rightarrow 2H_2 + O_2 \]
   
   a. One reactant in this process is _______.
   b. One product in this process is _______.

22. The color in the test tube changes from _______ to _______ when mercuric oxide is heated.

23. Which of the following is a physical change and which is a chemical change?
   - a. Hydrogen and oxygen produce water ____________.
   - b. Chlorine gas has a maximum solubility in water at 49°F.___________
   - c. Isopropyl alcohol forms water and carbon dioxide when burned.______________
   - d. Nitrogen gas is very compressible _____________

24. A ____________ is a solid product that can form in a liquid mixture.

Short Answers:

25. A copper penny is left outside for a long period of time and becomes green. Is this a chemical change or a physical change? How would you tell?
26. When iron is heated, it gives off a red glow. Is this a sign that a chemical change has taken place?
Answer Key

1. b
2. c
3. d
4. c
5. a
6. b
7. b
8. c
9. b
10. d
11. false
12. false
13. false
14. true
15. false
16. true
17. false
18. true
19. false
20. false
   a. H₂O
   b. O₂ or H₂
21. red, silver
   a. chemical
   b. physical
   c. chemical
   d. physical
22. precipitate
23. Chemical change. Can analyze the green material and find it is a chemical compound.
24. No, this is a physical process. The iron still has the same composition after it is cooled.
Chapter Test

Name___________________ Class________________ Date________

Multiple Choice:

1. One of the following is not a physical property of osmium
   a. density of 22.59 g/cm³
   b. boiling point of 5012°C
   c. forms osmium tetroxide with air
   d. silver metal

2. A reversible change is
   a. trimming a tree
   b. cooking a steak
   c. cutting up a potato
   d. making ice cubes

3. Which of the following is not true of gases?
   a. a hydrogen sample has a constant volume when the temperature changes
   b. oxygen can be compressed easily
   c. nitrogen molecules are very far apart in the gaseous state
   d. molecules of chlorine gas are constantly moving

4. Which of the following is a pure substance?
   a. air
   b. nitrogen
   c. iron alloy
   d. salt water

5. Only one of the following is a homogeneous mixture
   a. chocolate chip ice cream
   b. iced tea with lemon
   c. coffee with sugar
   d. root beer float

6. Sodium chloride forms sodium ions and chloride ions when dissolved in water. This process represents a
   a. physical change
   b. chemical change
   c. nuclear change
   d. elemental change

7. One of the following is not a characteristic of the compound NaHCO₃
   a. composed of four elements
   b. cannot be broken down into simpler substances
c. can undergo chemical changes
d. elements are combined in fixed proportions

8. The symbol Sn represents the element
   a. arsenic
   b. antimony
   c. mercury
   d. tin

9. The Latin word for potassium is
   a. natrium
   b. kalium
   c. plumbum
   d. stibium

10. A gas is produced when zinc reacts with
    a. hydrochloric acid
    b. sulfur
    c. mercuric oxide
    d. water

**True/False:**

11. _____ A chemical symbol for an element can have up to three letters.
12. _____ Water is fairly incompressible.
13. _____ A physical change does not involve a change in the identity of the sample.
14. _____ Mercury can be broken down into simpler components.
15. _____ Nitrogen at room temperature is very compressible.
16. _____ A homogeneous mixture consists of two or more phases.
17. _____ Iodine gas is formed from the reaction between potassium iodide and water.
18. _____ Distillation separates materials on the basis of boiling point.
19. _____ Distillation columns are cooled by compressed air.
20. _____ Boiling water is a chemical change.

**Matching:**

21. _____ the chemical symbol for iron.
22. _____ a form of matter with a definite volume, but an indefinite shape.
23. _____ separate gold dust from river water.
24. _____ a solid product that can form in a liquid mixture.

   a. liquid
   b. filtration
   c. precipitate
   d. Fe

**Fill in the blank:**

25. A _____ is formed when mercury is heated above its boiling point of 356.7°C.
26. The chemical symbol for gold is _____ and the Latin name is ______________.
27. Rust is the product of the reaction between iron and ____________.
28. Zinc sulfide is the product formed in the reaction between _____ and _____.
29. _____ is the measure of the amount of matter that an object contains.
30. A _____ forms when mercury is boiled.

Short Answers:

31. Can air exist as a pure substance? Explain your answer.
32. How would you separate a mixture of salt, olive oil, and water?

Answer Key

1. c
2. d
3. a
4. b
5. c
6. a
7. b
8. d
9. b
10. a
11. true
12. true
13. true
14. false
15. true
16. false
17. false
18. true
19. false
20. false
21. d
22. a
23. b
24. c
25. vapor
26. Au, aurum
27. air
28. zinc, sulfur
29. mass
30. new products
31. No, because the composition of air varies from place to place and depends on the conditions at a specific time.
32. The salt would dissolve in the water. Olive oil does not mix with water and will float to the top where it can be poured off. The water can then be boiled off (and condensed if need be) and the salt will be left behind.
CHAPTER 3

Measurements Assessments

Chapter Outline

3.1 THE INTERNATIONAL SYSTEM OF UNITS
3.2 UNIT CONVERSIONS
3.3 UNCERTAINTY IN MEASUREMENTS
3.4 MEASUREMENTS
Lesson Quiz

Name___________________ Class______________ Date________

Multiple Choice:

1. The SI symbol K stands for
   a. kilogram
   b. kilometer
   c. kelvin
   d. kilo

2. The SI symbol for the amount of a substance is
   a. mol
   b. mole
   c. m
   d. molar

3. 1500 megabytes is equivalent to
   a. $1.5 \times 10^3$ gigabytes
   b. $1.5 \times 10^2$ kilobytes
   c. $1.5 \times 10^{-2}$ gigabytes
   d. $1.5 \times 10^6$ kilobytes

4. The prefix hecto has an exponential factor of
   a. $1 \times 10^2$
   b. $1 \times 10^{-2}$
   c. $1 \times 10^4$
   d. $1 \times 10^{-4}$

5. It takes ___________ dekaliters to make 100 L.
   a. 1000
   b. 1
   c. 10
   d. 0.1

6. The Kelvin scale has ______ degrees between the melting point of ice and the boiling point of water.
   a. 125
   b. 150
   c. 75
   d. 100

7. The freezing point of water on the Kelvin scale is
   a. 273.15
   b. 373.15
8. The number 0.0015 can be represented by
   a. $1.5 \times 10^{-3}$
   b. $1.5 \times 10^{-2}$
   c. $1.5 \times 10^{-3}$
   d. $1.5 \times 10^{-2}$

9. The boiling point of water on the Celsius scale is
   a. 0°C
   b. 90°C
   c. 110°C
   d. 100°C

10. On which astronomical body will ten pounds of water weight less than it will on the Earth?
    a. Jupiter
    b. the moon
    c. the sun
    d. Arcturus

True/False:

11. _____ SI is the abbreviation for the Le Système International.
12. _____ A measurement includes both a number and a unit.
13. _____ The ampere is another unit for mass.
14. _____ There are 103 hg in 10.3 kg.
15. _____ Very large volumes may be conveniently expressed in microliters.
16. _____ There is no molecular motion at 0°C.
17. _____ Kinetic energy is the energy due to motion.
18. _____ The joule is the SI unit for energy.
19. _____ Particles of a cold material move faster than particles of a hot material.
20. _____ The Fahrenheit scale is commonly used for scientific work.

Fill in the Blank:

21. Give the decimal or scientific notation for the following:
    a. 10956. _____
    b. $67.3 \times 10^2$ _____
    c. 0.0045 _____
    d. $8.2 \times 10^{-2}$ _____

22. The freezing point of water on the Celsius scale is _____°C.
23. _____________ is a measure of the average _________ energy of the particles in matter.
24. Weight measures the effect of _________ on an object.
25. When two objects at different temperatures are brought into contact with one another, heat flows from the object at the _______ temperature to the object at the _______ temperature.
26. The Celsius temperature scale was developed by the ________ astronomer ________ ________.

Short Answers:

27. In a space capsule in outer space, why do objects float around?
28. Why is the SI system easier to use than the British system of units (feet, inches, pounds, ounces)?
3.1. The International System of Units

Answer Key

1. c
2. a
3. d
4. a
5. c
6. d
7. a
8. c
9. d
10. b
11. false
12. true
13. false
14. true
15. false
16. false
17. true
18. true
19. false
20. false
   a. \(1.0956 \times 10^4\)
   b. 6730
   c. \(4.5 \times 10^{-3}\)
   d. 0.082

21. 0°C
22. temperature, kinetic
23. gravity
24. higher, lower
25. Swedish, Anders Celsius
26. There is no gravitational pull to hold them in place.
27. The SI system is based on units of ten, while the British system has no common conversion factors when going between different sizes of units.
Lesson Quiz

Name___________________ Class________________ Date________

Multiple Choice:

1. The metric system uses powers of _______ to make conversions.
   a. 100
   b. 10
   c. 0.1
   d. 0.01

2. A conversion factor is a ratio of __________ measures
   a. opposite
   b. larger
   c. smaller
   d. equivalent

3. To convert from centimeters to meters, multiple the centimeter value by
   a. 10
   b. 0.1
   c. 100
   d. 0.01

4. Area is a unit derived from the base units
   a. \( \text{length} \times \text{mass} \)
   b. \( \text{width} \times \text{volume} \)
   c. \( \text{length} \times \text{volume} \)
   d. \( \text{length} \times \text{width} \)

5. The units for speed are
   a. m/s
   b. kg/s
   c. m/L
   d. m\(^2\)/s

6. Density is
   a. \( \text{mass} \times \text{volume} \)
   b. an intensive property
   c. \( \frac{\text{volume}}{\text{mass}} \)
   d. an extensive property

7. The SI unit for concentration is mol/L. What does “mol” represent?
   a. volume of material
   b. temperature of material
3.2. Unit Conversions

8. Dimensional analysis
   a. measures room size
   b. uses units to set up problems correctly
   c. defines SI units
   d. eliminates the need for a calculator

True/False:

9. _____ Helium is heavier than carbon dioxide.
10. _____ Ice will sink in 100% ethanol.
11. _____ The SI unit of force is the newton.
12. _____ One hour = 360 seconds.
13. _____ 4 cups/2 pints is a conversion factor.
14. _____ In using dimensional analysis, the units do not need to be included in the calculations.
15. _____ The SI unit for energy is the joule.
16. _____ The official SI unit for density is g/m³.

Fill in the Blank:

17. _____ has a density of 19.3 g/mL.
18. There are _________ mm in 12 cm.
19. If 4 cups = two pints and 4 cups = 1 quart, there are ________ pints in 1 quart.
20. A _______ unit is a unit that results from a mathematical combination of SI base units.

Perform the Following Calculations

21. There are 2.54 cm in one inch. I only have a centimeter ruler, but need to cut a 12 inch strip of duct tape. How much duct tape will I cut?
22. The lab instructor wants to prepare enough NaBr solution so that each of the 25 students will have 40 mL. He only has a two liter flask to store the solution in. Will there be enough room for all the solution?
23. A runner decides to participate in a 7.2 km run, but only knows miles (he did not pay attention to the chemistry lecture on units). He does learn that there are 1.6 km in a mile. How far (in miles) will he be running?
24. How many µL are there in 0.0245 mL?
25. Gold is selling for $37.33/gram. I have 45.7 grams of pure gold. How much money will I get when I sell it?
26. Gold has a density of 19.3 g/mL. I have a metal sample that I was told was pure gold. It has a mass of 397.2 grams and a volume of 33.1 mL. Is this a pure sample of gold?
27. A swimming pool has the following dimensions: 20 m × 10 m × 5 m. What is the volume of the pool?
28. An old folk song starts “It’s a long hard road from Lynchburg to Danville ...”. The distance between these two towns in southern Virginia is 70.3 miles. If a car travels at 55 miles/hour, how long will it take to get from one town to the other?
29. Why is a golf ball heavier than a tennis ball?
30. Helium gas has a density of 0.166 g/L at 20°C while radon gas has a density of 9.23 g/L at the same temperature. What can we infer about the relative sizes of the two types of atoms?

Answer Key

1. b
2. d  
3. c  
4. d  
5. a  
6. b  
7. c  
8. b  
9. false  
10. true  
11. true  
12. false  
13. true  
14. false  
15. true  
16. false  
17. gold  
18. 120  
19. 2  
20. derived  
21. \((12 \text{ inches})(2.54 \text{ cm/in}) = 30.5 \text{ cm}\)  
22. \((25 \text{ students})(40 \text{ mL/student}) = 1000 \text{ mL}\) solution needed. Yes, there is enough room in the flask.  
23. \((7.2 \text{ cm})(\frac{1 \text{ mile}}{1.6 \text{ km}}) = 4.5 \text{ miles}\)  
24. \((0.0245 \text{ mL})(1000 \mu \text{L/mL}) = 24.5 \mu \text{L}\)  
25. \((37.33 \text{ g})(45.7 \text{ g}) = 1705.98\)  
26. density of unknown material = \(\frac{(397.2 \text{ g})}{(33.1 \text{ mL})} = 12.0 \text{ g/mL}\) – definitely not gold  
27. \(20 \text{ m} \times 10 \text{ m} \times 5 \text{ m} = 1000 \text{ m}^3\) volume  
28. \((70.3 \text{ miles})(\frac{1 \text{ hour}}{55 \text{ miles}}) = 1.3 \text{ hours}\)  
29. A golf ball has a smaller volume, but it is solid. The tennis ball has an empty interior. Therefore, the golf ball has a higher density.  
30. Under similar conditions, we can assume the same number of gas atoms in each case. For a given volume of gas, helium has a lower density and will have a lower mass. Since there are the same number of molecules, the helium atoms must be smaller than the radon atoms.
**Lesson Quiz**

Use the diagrams below to answer the first three questions. Each X represents data points and the O indicates the accepted value.

1. Data set one is
   a. accurate but not precise
   b. precise but not accurate
   c. both accurate and precise
   d. neither accurate nor precise

2. Data set two is
   a. accurate but not precise
   b. precise but not accurate
   c. both accurate and precise
   d. neither accurate nor precise

3. Data set three is
   a. accurate but not precise
   b. precise but not accurate
   c. both accurate and precise
   d. neither accurate nor precise

4. The experimental value of a measurement is
   a. voted on by the lab team
   b. obtained from the instructor
   c. looked up on the internet
   d. measured during the experiment

5. The percent error is
   a. an absolute value
b. the ratio of experimental value to accepted value
c. the ratio of accepted value to experimental value
d. always a whole number

6. The amount of uncertainty in a measurement depends upon
   a. the length of the ruler
   b. operator error only
   c. calibration of the balance
   d. operator error and measurement equipment

7. The inventory report shows 23 graduated cylinder in the stockroom. How many significant figures are in this number?
   a. one
   b. two
   c. none
   d. infinite

8. Insignificant figures
   a. are always ignored
   b. help locate decimal points
   c. are never reported
   d. indicate accuracy

9. The value 0.008140 has ____ significant figures.
   a. one
   b. two
   c. three
   d. four

10. For addition and subtraction, it is the position of the _____ that influences the rounding of the value.
    a. decimal point
    b. leading zero
    c. first digit
    d. final digit

True/False:
11. _____ Precision is a measure of how close the values are to the accepted value.
12. _____ The percent error cannot be calculated if the experimental value and the accepted value are the same.
13. _____ When calculating percent error, the accepted value is in the denominator.
14. _____ The significant figures in a measurement consist of all the certain digits in that measurement plus one uncertain or estimated digit.
15. _____ All nonzero digits in a number are significant.
16. _____ When 1.023 is divided by 0.447, the answer will be rounded to four significant figures.
17. _____ In rounding numbers in addition and subtraction problems, use the number of decimal places as a guide in deciding how to round.
18. _____ Zeros that appear between other nonzero digits are always significant.

Fill in the Blank:

19. Fill in the following table:
3.3. Uncertainty in Measurements

**Table 3.1:**

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Accepted Value</th>
<th>Experimental Value</th>
<th>Percent Error</th>
</tr>
</thead>
<tbody>
<tr>
<td>a.</td>
<td>14.85</td>
<td>12.99</td>
<td></td>
</tr>
<tr>
<td>b.</td>
<td>14.85</td>
<td>16.32</td>
<td></td>
</tr>
<tr>
<td>c.</td>
<td>14.85</td>
<td>+2.7</td>
<td></td>
</tr>
</tbody>
</table>

20. Fill in the following table:

**Table 3.2:**

<table>
<thead>
<tr>
<th>Number</th>
<th>How Many Significant Figures?</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. 7.2 × 10⁻⁴</td>
<td></td>
</tr>
<tr>
<td>b. 33.709</td>
<td></td>
</tr>
<tr>
<td>c. 1408</td>
<td></td>
</tr>
<tr>
<td>d. 2.69 × 10⁴</td>
<td></td>
</tr>
</tbody>
</table>

21. Give the answer (with the correct number of significant figures) for the following calculations:

a. \( 67 \times 23.12 = \)

b. \( 867 + 23.4 = \)

c. \( \frac{805}{35} = \)

d. \( 296.4 - 39.1 = \)

**Short Answers:**

22. Why would we have a rounding rule (below 5, drop; above 5 round up)?

23. How can we be confident that the density of gold is really 19.3 g/mL?

**Answer Key**

1. d
2. c
3. b
4. d
5. b
6. d
7. d
8. b
9. d
10. a
11. false
12. false
13. true
14. true
15. true
16. false
17. true
18. true
Table 3.3:

<table>
<thead>
<tr>
<th>Experiment</th>
<th>Accepted Value</th>
<th>Experimental Value</th>
<th>Percent Error</th>
</tr>
</thead>
<tbody>
<tr>
<td>a.</td>
<td>14.85</td>
<td>12.99</td>
<td>12.5%</td>
</tr>
<tr>
<td>b.</td>
<td>14.85</td>
<td>16.32</td>
<td>9.9%</td>
</tr>
<tr>
<td>c.</td>
<td>14.85</td>
<td>15.25</td>
<td>+2.7</td>
</tr>
</tbody>
</table>

Table 3.4:

<table>
<thead>
<tr>
<th>Number</th>
<th>How Many Significant Figures?</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. $7.2 \times 10^{-4}$</td>
<td>two</td>
</tr>
<tr>
<td>b. 33.709</td>
<td>five</td>
</tr>
<tr>
<td>c. 1408</td>
<td>four</td>
</tr>
<tr>
<td>d. $2.69 \times 10^4$</td>
<td>three</td>
</tr>
</tbody>
</table>

21.  1. 1500  
      2. 890  
      3. 23  
      4. 257

22. If several numbers are involved, the difference will average out and cancel one another. If all numbers were rounded only one way (either higher or lower), a gradual error would be introduced into the data set.

23. Repeated measurements by highly qualified persons using very accurate equipment will give a reliable value
Chapter Test

Name___________________ Class______________ Date________

Multiple Choice:

Three different technicians measured the amount of silver in an ore sample. The data for each technician are listed below. The accepted value as determined by a certified laboratory is 14.07 grams. Use this data to answer the first three multiple choice questions.

**Table 3.5:**

<table>
<thead>
<tr>
<th>Analysis</th>
<th>Technician One</th>
<th>Technician Two</th>
<th>Technician Three</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>12.98</td>
<td>13.98</td>
<td>11.45</td>
</tr>
<tr>
<td>2</td>
<td>17.62</td>
<td>14.22</td>
<td>11.63</td>
</tr>
<tr>
<td>3</td>
<td>10.51</td>
<td>13.87</td>
<td>12.09</td>
</tr>
<tr>
<td>4</td>
<td>7.35</td>
<td>14.15</td>
<td>11.84</td>
</tr>
</tbody>
</table>

1. Data from technician one is
   a. accurate but not precise
   b. precise but not accurate
   c. both accurate and precise
   d. neither accurate nor precise

2. Data from technician two is
   a. accurate but not precise
   b. precise but not accurate
   c. both accurate and precise
   d. neither accurate nor precise

3. Data from technician three is
   a. accurate but not precise
   b. precise but not accurate
   c. both accurate and precise
   d. neither accurate nor precise

4. The SI unit for mass is
   a. g
   b. kg
   c. mol
   d. K

5. It takes _________ L to make 5 kL.
   a. 500
   b. 5
6. Density is
   a. \( \frac{mass}{volume} \)
   b. \( \frac{mass}{mass} \)
   c. \( mass \times volume \)
   d. \( \frac{(mass)^2}{volume} \)

7. The parameters that determine the joule are
   a. \( force \times length \)
   b. \( force \times mass \)
   c. \( force \times volume \)
   d. \( force \times area \)

8. The prefix milli has an exponential factor of
   a. \( 1 \times 10^{-9} \)
   b. \( 1 \times 10^{-6} \)
   c. \( 1 \times 10^{-3} \)
   d. \( 1 \times 10^{-12} \)

9. To convert from kilograms to grams, multiply the kilogram value by
   a. 10
   b. \( 1 \times 10^2 \)
   c. \( 1 \times 10^2 \)
   d. \( 1 \times 10^4 \)

10. The value 1.070 has ___ significant figures
    a. two
    b. four
    c. one
    d. three

True/False:
11. _____ N is another unit for mass
12. _____ Very small volumes may conveniently be expressed as \( \mu L \).
13. _____ SI is the abbreviation for Le Système International d’Unités.
14. _____ A measurement includes only a number.
15. _____ 4 quarts/1 gallon is a conversion factor.
16. _____ Carbon dioxide has a higher density than radon.
17. _____ Temperature is an indicator of particle kinetic energy.
18. _____ All gases have a density less than that of water.
19. _____ The Swiss astronomer Anders Celsius developed the Celsius temperature scale.
20. _____ Precision is a measure of how close the experimental values are to one another.

Fill in the Blank:
21. The boiling point of water on the Kelvin scale is ________ °K.
22. There are _______ mL in 1098 \( \mu L \).
23. If there are 2.54 cm in an inch and 12 inches in a foot, how many cm are in a foot?
24. What is the area in square centimeters of a small pond that measures 3.6 m \( \times \) 7.2 m?
25. The speed of light is \(3.0 \times 10^8 \text{ m/s}\). The moon is 384,400 km from the earth. How long does it take light reflected by the moon to reach us?

26. A small cruise ship contains 354,000 gallons of fuel. If the ship sails without stopping and burns 936 gallons of fuel per hour, will it be able to sail from London to New York (3459 miles) at a speed of 20 miles/hour without running out of fuel?

**Short Answers:**

27. Archimedes was asked to determine whether the gold in the king’s crown was actually pure gold. How would you help Archimedes solve this problem using nothing more than a set of scales and a basin of water?

28. The United States “officially” uses the British system of pounds, inches, and gallons (among others). Would there be any advantages to changing all our measurements to the metric system like the rest of the world? Explain your answer.

---

**Answer Key**

1. d  
2. c  
3. b  
4. b  
5. c  
6. a  
7. a  
8. c  
9. c  
10. d  
11. false  
12. true  
13. true  
14. false  
15. true  
16. false  
17. true  
18. false  
19. false  
20. true  
21. 371  
22. 1.098  
23. \((2.54 \text{ cm/inch})(\frac{12 \text{ inches}}{\text{ foot}}) = 30.48 \text{ cm/foot}\)  
24. \((3.6 \text{ m})(100 \text{ cm/m}) = 360 \text{ cm}\)  
   \((7.2 \text{ m})(100 \text{ cm/m}) = 720 \text{ cm}\)  
   \((360 \text{ cm})(720 \text{ cm}) = 2.59 \times 10^5 \text{ cm}^2\)  
   - \((3.84 \times 10^5 \text{ km})(1 \times 10^3 \text{ m/km}) = 3.84 \times 10^8 \text{ m}\)  
25. \(\frac{3.84 \times 10^8 \text{ m}}{30 \times 10^6 \text{ m/s}} = 1.28 \text{ seconds}\)  
26. \(\frac{\text{354,000 gallons}}{936 \text{ gallons/hr}} = \text{enough fuel for 378 hours of sailing}\)  
   \(\frac{3459 \text{ miles}}{20 \text{ miles/hr}} = 173 \text{ hours needed for sailing}\)  
   There is more than enough fuel for the voyage.
27. Weigh the crown and determine the volume by measuring the water displaced. Do the same with a known sample of pure gold. Calculate the densities of both materials and see if they match.
28. Yes, there would be advantages. Everyone would only need one set of measuring devices for all measurements. Manufacturing plants would only need to make items in one set of units, not two.
### Chapter Outline

<table>
<thead>
<tr>
<th>Section</th>
<th>Title</th>
</tr>
</thead>
<tbody>
<tr>
<td>4.1</td>
<td>Atoms</td>
</tr>
<tr>
<td>4.2</td>
<td>The Nuclear Model of the Atom</td>
</tr>
<tr>
<td>4.3</td>
<td>Isotopes and Atomic Mass</td>
</tr>
<tr>
<td>4.4</td>
<td>Atomic Structure</td>
</tr>
</tbody>
</table>
Lesson Quiz

Name___________________ Class________________ Date________

Multiple Choice:

1. Democritus believed the atom was
   a. indivisible
   b. made up of smaller particles
   c. universally accepted
   d. experimentally verified

2. Dalton believed that atoms could be
   a. broken into parts
   b. rearranged to create other atoms
   c. combined to form compounds
   d. used to form elements

3. By the 1700s scientists knew about
   a. neutrons
   b. elements
   c. electrons
   d. protons

4. Samples of a specific compound from different sources obey the law of
   a. multiple proportions
   b. simple proportions
   c. definite proportions
   d. ideal proportions

5. Our idea of the atom
   a. has not changed since Democritus
   b. has been fairly constant for 200 years
   c. is essentially what John Dalton described
   d. is very different from what Dalton described

6. Chemical reactions change the ______ of atoms
   a. composition
   b. arrangement
   c. size
   d. shape

7. Methane (one C, 4 H) and ethane (2 C, 6H) illustrate the law of
   a. definite proportions
   b. multiple proportions
4.1. Atoms

8. The law of conservation of mass says that the mass of products in a reaction
   a. is greater than the mass of reactants
   b. is less than the mass of reactants
   c. is equal to the mass of reactants
   d. all of the above

9. Atoms combine in simple ___________ to form compounds.
   a. proportions
   b. fractional ratios
   c. whole number ratios
   d. variable ratios

10. A rearrangement of atoms could be called
    a. chemical reaction
    b. nuclear fission
    c. transmutation
    d. elemental partition

True/False:

11. _____ Carbon and oxygen can only combine to form one compound.
12. _____ The mass of reactants must equal the mass of products.
13. _____ John Dalton was the first person to propose a detailed structure for the atom.
14. _____ Mass ratio must be a whole-number ratio.
15. _____ Chemical reactions can change the identity of atoms.
16. _____ All samples of pure water contain the same ratio of H to O.
17. _____ Dalton predicted the idea of the neutron.
18. _____ The ideas of Democritus were rejected because there was no way to prove them.
19. _____ Dalton’s ideas about the atom were based (in part) on the law of definite proportions.
20. _____ Dalton’s theory was proposed in 1743.

Fill in the Blank:

21. The Greek word “atomos” means ______________.
22. The approach of Democritus was ____________, not scientific.
23. Formation of 46 grams of product from 46 grams of reactant illustrates the law of ________ of mass.
24. In the 1790s, a greater emphasis began to be placed on the ___________ analysis of chemical reactions.
25. Mass cannot be __________ or __________ during a chemical reaction.
26. A given chemical compound always contains the same elements in the exact same ________ by mass.
27. An ________ is the smallest particle of an element that retains the properties of that element.
28. The sizes of atoms of the same element are ________.
29. Lead forms two compounds with oxygen. One compound contains 2.98 g of lead and 0.461 g of oxygen. The other contains 9.89 g of lead and 0.763 g of oxygen. For a given mass of oxygen, what is the lowest whole-number mass ratio of lead in the two compounds?

Short Answers:

30. What is a model? How does a model of the atom help understand a real atom?
31. A certain reaction uses 34.6 grams of reactants and forms 19.3 grams of product A plus a gas. Does this disprove the law of conservation of mass? Explain your answer.
Answer Key

1. a
2. c
3. b
4. c
5. d
6. b
7. b
8. c
9. c
10. a
11. false
12. true
13. false
14. true
15. false
16. true
17. false
18. true
19. true
20. false
21. indivisible
22. philosophical
23. conservation
24. quantitative
25. created, destroyed
26. ratio
27. atom
28. identical
29. compound one: 2.98 g lead/0.461 g O = 6.46:1
   compound two: 9.89 g lead/0.763 g O = 13.0:1
   mass ratio of Pb 6.46/13.01 = 0.5:0.2 = 1:2
30. A model gives a picture or an idea of the real thing. The model helps us understand and develop new questions.
31. No, the remainder of the product mass is in the gas which was not weighed.
Lesson Quiz

Name___________________ Class______________ Date________

Multiple Choice:

1. Dalton had a better idea of the atom than Democritus because
   a. Dalton was smarter
   b. Dalton had studied more science
   c. Democritus did no experiments
   d. Dalton had a better lab

2. A cathode ray tube emits
   a. protons and electrons
   b. electrons only
   c. protons only
   d. neither protons or electrons

3. Thomson showed the cathode ray had mass by using a
   a. electronic balance
   b. magnet
   c. electric current
   d. paddle wheel

4. Rutherford used ______ foil to study atomic structure.
   a. aluminum
   b. lead
   c. gold
   d. platinum

5. Millikan’s ______ drop experiments helped determine some properties of the electron.
   a. water
   b. oil
   c. octane
   d. propane

6. A neutron has a charge of
   a. +1
   b. -1
   c. 0
   d. -2

7. What did Eugene Goldstein first call protons?
   a. channel rays
   b. canal rays
8. In the “plum pudding” model of the atom, the electrons were
   a. mixed in the nucleus with the protons
   b. orbiting around the protons
   c. layered on the surface of the proton mass
   d. surrounded by the protons

9. The plum pudding model was proposed by
   a. Eugene Goldstein
   b. Robert Millikan
   c. Ernest Rutherford
   d. J.J. Thomson

10. Rutherford’s atomic model became known as the ______ model.
    a. neutron
    b. nucleus
    c. nuclear
    d. noodle

True/False:

11. _____ Dalton assumed that atoms could be further divided.
12. _____ The electron and the neutron have essentially the same mass.
13. _____ The Rutherford model of the atom came after the Thomson model.
14. _____ The charge on the neutron is zero.
15. _____ The electron has a mass 1/1840 of that of the proton.
16. _____ Millikan originally called the electron a corpuscle.
17. _____ Goldstein discovered the proton.
18. _____ Most alpha particles were not deflected by the gold foil.
19. _____ Rutherford concluded that the electrons deflected the alpha particles.
20. _____ Electrical charges are carried by particles of matter.

Fill in the Blank:

21. Millikan’s studies measured the _____ of an electron.
22. A cathode ray tube has _____ at either end of the tube.
23. Cathode rays were deflected __________ a negatively charged metal plate and ________ a positively charged plate.
24. The _____ is a positively charged subatomic particle that is present in all atoms.
25. An alpha particle is about _____ times the size of a hydrogen atom.
26. A small number of alphas particles bounced off the gold foil at __________ angles.
27. The unique _____ and ________ of subatomic particles gives each element its specific identity.
28. The _______ was the first subatomic particle discovered.
29. The cathode ray travels from the ______ to the __________.
30. __________ charges attract one another.

Short Answer:

31. How did the discovery of the electron lead to other discoveries about the atom?
32. How would Thomson have interpreted the gold foil experiment to support his “plum pudding” model?
4.2. The Nuclear Model of the Atom

Answer Key

1. c
2. b
3. d
4. c
5. b
6. c
7. b
8. a
9. d
10. c
11. false
12. false
13. true
14. true
15. true
16. false
17. true
18. true
19. false
20. true
21. charge
22. electrodes
23. away from, toward
24. proton
25. four
26. large
27. number, arrangement
28. electrons
29. cathode, anode
30. opposite
31. By finding a negatively charged particle, scientists had to assume there was a positively charged particle also present.
32. The plum pudding was just too thin to block the alpha particles.
Lesson Quiz

Name___________________ Class________________ Date________

Multiple Choice:

1. The number of protons represents the
   a. atomic mass
   b. atomic number
   c. atomic weight
   d. atomic isotope

2. The number of electrons in an element equals
   a. the number of neutrons
   b. the number of protons
   c. the number of protons + neutrons
   d. the number of protons – neutrons

3. An element has an atomic number of 25. That element is
   a. Na
   b. Sr
   c. Mn
   d. Cr

4. One proper way to indicate an isotope of Si with an atomic mass of 28 is
   a. 28-Si
   b. Si-28
   c. \(^{28}\text{Si}\)
   d. \(^{28}\text{Si}\)

5. How many neutrons are in the nucleus of \(^{96}_{42}\text{Mo}\)?
   a. 53
   b. 42
   c. 48
   d. 54

6. Two isotopes of an element differ in the number of
   a. neutrons
   b. protons
   c. electrons
   d. muons

7. The carbon-12 nuclide is used to define the
   a. atomic mass
   b. atomic weight
4.3. Isotopes and Atomic Mass

8. The most abundant isotope of oxygen is
   a. O-16
   b. O-17
   c. O-18
   d. O-19

9. Average atomic mass is calculated from
   a. atomic mass and atomic number
   b. average of isotope masses
   c. atomic masses and abundance of isotopes
   d. relative abundance of isotopes

10. A mass spectrometer can determine
    a. proton mass
    b. neutron mass
    c. electron mass
    d. atomic mass

True/False:

11. _____ Most elements have only two isotopes.
12. _____ Isotopic composition affects chemical reactivity.
13. _____ “Nuclide” refers to the nucleus of a given isotope of an element.
14. _____ The symbol $^{82}_{\text{Pb}}$ indicates an element with an atomic mass of 82.
15. _____ Ge-73 is another way of showing atomic mass.
16. _____ The isotope $^{101}_{44}\text{Ru}$ has 57 neutrons.
17. _____ The isotope in problem 16 has 47 electrons.
18. _____ Atoms are electrically neutral.
19. _____ The presence of isotopes supports Dalton’ atomic theory.
20. _____ The number of neutrons = mass number – atomic number.

Fill in the Blank:

<table>
<thead>
<tr>
<th>Table 4.1:</th>
</tr>
</thead>
<tbody>
<tr>
<td>Element</td>
</tr>
<tr>
<td>Indium</td>
</tr>
<tr>
<td>Yttrium</td>
</tr>
<tr>
<td>Sulfur</td>
</tr>
<tr>
<td>Rubidium</td>
</tr>
<tr>
<td>Argon</td>
</tr>
</tbody>
</table>

22. Calculate the average atomic mass of the following:
   In-113 4.29% and I-115 95.7%

23. Neutrons do not contribute to chemical reactivity and new compound formation. What changes would we have to make in writing chemical symbols if neutrons somehow changed in the course of a reaction?
Chapter 4. Atomic Structure Assessments

Answer Key

1. b
2. b
3. c
4. b
5. d
6. a
7. d
8. a
9. c
10. d
11. false
12. false
13. true
14. true
15. false
16. true
17. false
18. true
19. false
20. true

Table 4.2:

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Protons</th>
<th>Neutrons</th>
<th>Electrons</th>
<th>Atomic Number</th>
<th>Atomic Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Indium</td>
<td>In</td>
<td>49</td>
<td>65</td>
<td>49</td>
<td>49</td>
<td>114</td>
</tr>
<tr>
<td>Yttrium</td>
<td>Y</td>
<td>39</td>
<td>50</td>
<td>39</td>
<td>39</td>
<td>89</td>
</tr>
<tr>
<td>Sulfur</td>
<td>S</td>
<td>16</td>
<td>16</td>
<td>16</td>
<td>16</td>
<td>32</td>
</tr>
<tr>
<td>Rubidium</td>
<td>Rb</td>
<td>37</td>
<td>48</td>
<td>37</td>
<td>37</td>
<td>85</td>
</tr>
<tr>
<td>Argon</td>
<td>Ar</td>
<td>18</td>
<td>22</td>
<td>18</td>
<td>18</td>
<td>40</td>
</tr>
</tbody>
</table>

22. 

\[
\text{average atomic mass} = (113)(0.0429) + (115)(0.957) \\
= 4.85 + 110.1 \\
= 114.95
\]

23. We would have to find new symbols to show those changes. Currently, the symbol for an element does not change when it becomes part of a new compound, but we would need new symbols to show neutron changes.
Chapter Test

Name___________________ Class________________ Date________

Multiple Choice:

1. Which of the following reactions illustrates the law of conservation of mass?
   a. $43.2 \text{ g A} + 17.9 \text{ g B} \rightarrow 30.1 \text{ g C} + 31.0 \text{ g D}$
   b. $68.2 \text{ g A} \rightarrow 27.1 \text{ g B} + 43.5 \text{ g C}$
   c. $12.3 \text{ g A} + 9.6 \text{ g B} \rightarrow 23.9 \text{ g C}$
   d. $18.7 \text{ g A} + 22.4 \text{ g B} \rightarrow 26.3 \text{ g C} + 21.6 \text{ g D}$

2. Rutherford used gold foil to study
   a. nuclear particles
   b. atomic structure
   c. neutron composition
   d. atomic masses

3. The oil drop experiment was performed by
   a. Thomson
   b. Goldstein
   c. Millikan
   d. Dalton

4. Canal rays was the first name given to
   a. neutrons
   b. electrons
   c. protons
   d. gluons

5. The nuclear model of the atom replaced the __________ model.
   a. plum pudding
   b. cherry pie
   c. fudge sundae
   d. plum tart

6. The atomic mass equals
   a. the number of protons
   b. the number of neutrons
   c. the number of protons + neutrons
   d. the number of electrons + neutrons

7. The symbol $^{108}_{47}\text{Ag}$ represents an isotope with
   a. atomic number 108 and atomic mass 155
   b. atomic number 108
c. atomic number 47 and atomic mass 61
d. atomic number 47 and atomic mass 108

8. Which of the following was not part of Democritus’ model of the atom?
   a. basic unit of matter
   b. indivisible
   c. smallest unit of matter
   d. reacted to form molecules

9. A chemical reaction is
   a. a rearrangement of atoms
   b. a dividing of atoms
   c. a creation of new atoms
   d. a disappearance of atoms

10. The cathode ray tube helped in the discovery of
    a. neutrons
    b. electrons
    c. atoms
    d. nuclides

True/False:

11. _____ Carbon and oxygen can form ten different compounds.
12. _____ Water has the same H:O ratio on Mars as it does on Earth.
13. _____ Millikan discovered the electron using an oil drop apparatus.
14. _____ The proton is 1840 times as heavy as the electron.
15. _____ The proton is much heavier than the neutron.
16. _____ The isotope $^{178}_{72}Hf$ has 106 neutrons.
17. _____ Cs-132 has 132 neutrons.
18. _____ Neutrons play an important role in chemical reactivity.
19. _____ The nucleus of an atom is electrically positive.
20. _____ The ideas of Democritus became popular very quickly.

Fill in the Blank:

21. The chemical and ______ properties of compounds are different than the properties of the __________ from which they were formed.
22. Mass is ______ conserved in chemical reactions.
23. Two carbon-oxygen compounds are carbon dioxide and carbon ______.
24. The three fundamental particles are called the __________, the __________, and the ______.

**Table 4.3:**

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Protons</th>
<th>Neutrons</th>
<th>Electrons</th>
<th>Atomic Number</th>
<th>Atomic Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>cesium</td>
<td>$^{133}_{55}Cs$</td>
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<tr>
<td>iridium</td>
<td>Ir</td>
<td>77</td>
<td></td>
<td></td>
<td></td>
<td>192</td>
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<tr>
<td>arsenic</td>
<td>As</td>
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<td>41</td>
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<td>74</td>
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<tr>
<td>palladium</td>
<td>Pd</td>
<td></td>
<td>46</td>
<td></td>
<td></td>
<td>106</td>
</tr>
</tbody>
</table>

26. Carbon and hydrogen form many compounds. Methane has 1 carbon and 4 hydrogens while ethane had 2
4.4. Atomic Structure

4.4.1. Atomic Structure

There are two isotopes of boron. Boron-10 has a mass of 10.012937 amu and makes up 19.9% of the total. Boron-11 has a mass of 11.009305 amu and comprises 80.1% of the total. Calculate the average atomic mass of boron.

**Answer Key**

1. a  
2. b  
3. c  
4. c  
5. a  
6. c  
7. d  
8. d  
9. a  
10. b  
11. false  
12. true  
13. false  
14. true  
15. false  
16. true  
17. false  
18. false  
19. true  
20. false  
21. physical, elements  
22. always  
23. monoxide  
24. proton, neutron, electron

**Table 4.4:**

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Protons</th>
<th>Neutrons</th>
<th>Electrons</th>
<th>Atomic Number</th>
<th>Atomic Mass</th>
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<td>46</td>
<td>60</td>
<td>46</td>
<td>46</td>
<td>106</td>
</tr>
</tbody>
</table>

26. Methane: \( \frac{4}{12} = 0.33 \)  
   Ethane: \( \frac{6}{15} = 0.4 \)  
   Ratio: \( \frac{6}{15} = 1.32 : 1 \)

27. Average atomic mass = \( (10.02)(0.199) + (11.01)(0.801) = 10.81 \) amu
Chapter 5: Electrons in Atoms Assessments

Chapter Outline

5.1 Light
5.2 The Quantum Mechanical Model
5.3 Electron Arrangement in Atoms
5.4 Electrons in Atoms
Lesson Quiz

Name___________________ Class________________ Date________

Multiple Choice:

1. The visible spectrum of light spans the range of approximately
   a. 300-500 nm
   b. 400-600 nm
   c. 300–700 nm
   d. 400-700 nm

2. Energy levels of electrons in an atom are
   a. continuous
   b. discrete
   c. random
   d. variable

3. A low-frequency wave has a ______ wavelength.
   a. long
   b. short
   c. variable
   d. high

4. The units of frequency are
   a. meters/second
   b. cycles/meter
   c. cycles/second
   d. meters/minute

5. The Brackett series of hydrogen emission lines represent drops in energy from higher levels to the _______-
   _ level.
   a. \( n = 1 \)
   b. \( n = 2 \)
   c. \( n = 3 \)
   d. \( n = 4 \)

6. The emission spectrum for helium contains _____ lines than the hydrogen spectrum.
   a. more
   b. the same
   c. less
   d. variable

7. The photoelectric effect shows that light
   a. is composed of waves
b. can be divided into discrete wavelengths  
c. has properties of a particle  
d. is continuous

8. The photoelectric effect was first demonstrated by  
   a. Albert Einstein  
   b. Robert Millikan  
   c. John Dalton  
   d. J.J. Thomson

9. The amplitude of a wave is its  
   a. frequency  
   b. height  
   c. wavelength  
   d. velocity

10. Frequency is represent by the Greek letter  
    a. η  
    b. σ  
    c. ν  
    d. λ

True/False:

11. _____ The wavelengths of infrared light are longer than those of visible light.  
12. _____ An electron and a photon are of equal mass.  
13. _____ Electrons move to lower energy levels as they emit energy.  
14. _____ Two of the hydrogen emission lines are in the infrared region.  
15. _____ The threshold frequency is the minimum frequency of light that will eject electrons from a surface.  
16. _____ The Bohr model of the atom explains the emission lines of iron.  
17. _____ Electromagnetic radiation exhibits wavelike behavior.  
18. _____ Light from a heated body is emitted at all wavelengths.  
19. _____ Light as a particle helps explain some experimental results.  
20. _____ Classical physics could completely explain the photoelectric effect.

Fill in the Blank:

21. The wavelengths of visible light are ______ than those of infrared light.  
22. A ______ is the minimum amount of energy that can be gained or lost by an electron.  
23. _______ cells are commonly found in common devices such as calculators.  
24. One hertz equals ______________.  
25. The ________ ________ for an atom is the lowest energy state for that atom.  
26. An atom is in an _______ state when its potential energy is higher than that of the ground state.  
27. An ________ _________ spectrum is seen when light emitted from an atom passes through a prism.  
28. The change in energy when an electron makes the transition from one energy level to another can be calculated using the term ______.  
29. X-rays have __________ energy levels than microwaves.

Calculate:

30. What is the frequency of light that has a wavelength of $6 \times 10^{-10}$ m?  
31. What is the frequency of light that has an energy of $3.71 \times 10^{-17}$ J?
5.1. Light

Short Answer:

32. Why does the gas in a gas discharge tube need to be at low pressure in order for the gas to emit light when an electric current is passed through the tube?

\[ c = \frac{\lambda}{\nu} \]

\[ \lambda = 3 \times 10^8 \text{ m/sec} \]

\[ \nu = \frac{6 \times 10^{-10}}{m} \]

\[ = 5 \times 10^{17} \text{ Hz} \]
31. 

\[ E = \frac{h}{\hbar} \]
\[ = \frac{3.71 \times 10^{-17} \text{ J}}{6.626 \times 10^{-34} \text{ J} \cdot \text{sec}} \]
\[ = 5.6 \times 10^{18} \text{ Hz} \]

32. At higher pressures, gas molecules could collide with one another and absorb energy from electrons.
5.2 The Quantum Mechanical Model

Lesson Quiz

Name___________________ Class______________ Date________

Multiple Choice:

1. The study of motion in large objects is called
   a. traditional mechanics
   b. quantum mechanics
   c. classical mechanics
   d. quantized mechanics

2. According to the Heisenberg Uncertainty Principle, the velocity and position of which of the following can never be measured very accurately?
   a. a traveling missile
   b. an electron in motion
   c. a hockey puck shot at the goal
   d. atomic emission lines

3. An electron cloud
   a. shows exactly where the electron is located
   b. gives a 75% probability of where the electron is located
   c. gives a 90% probability of where the electron is located
   d. gives a 65% probability of where the electron is located

4. Each electron can be described by _____ quantum numbers.
   a. two
   b. four
   c. six
   d. eight

5. The highest allowable sublevel for \( n = 3 \) is
   a. s
   b. f
   c. d
   d. p

6. The spin quantum number has _____ values.
   a. two
   b. three
   c. one
   d. four

7. The units of mass in the de Broglie wave equation are
   a. \( \mu g \)
8. An increase in the quantum number \( n \) means _______ in the distance of the electron from the nucleus.
   a. an increase
   b. no change
   c. a decrease
   d. a variable response

9. If the principal quantum number \( n = 4 \), the angular quantum number can have values for zero to
   a. 2
   b. 4
   c. 1
   d. 3

10. The magnetic quantum number tells us the
    a. shape of the orbital
    b. orientation of the orbital around the nucleus
    c. distance of the orbital from the nucleus
    d. spin of the orbital

True/False:

11. _____ The Bohr model of the atom only explains the behavior of the hydrogen electron.
12. _____ Electrons in atoms do not have defined energy levels.
13. _____ It is not possible to measure simultaneously the exact velocity and location of a jet plane traveling at 645 miles/hour.
14. _____ The Schrödinger wave equation predicts the probability of finding an electron in a given area of space.
15. _____ An orbital represents a two-dimensional area of space.
16. _____ The spin quantum number can have up to four values.
17. _____ The s orbital has a spherical shape.
18. _____ The principal quantum number designates the principal energy level occupied by the electron.
19. _____ The term for the discrete energy levels of an electron is quantal.
20. _____ Two electrons in the same orbital cannot have the same spin quantum number.

Fill in the Blank:

21. The following image is a depiction of a ____ orbital.

22. A d orbital electron can assume any one of _____ different orientations.
23. If the angular momentum quantum number has an \( l \) of 0, that designates an ___ orbital.
24. The study of the motions of subatomic particles is called ______ ________.
25. Solutions to the Schrödinger equation are called _______ functions.
26. Each electron can be described by ____ quantum numbers.
27. A maximum of ____ electrons can be found in p orbitals.
28. The equation that described the wave nature of any particle was developed by ______ _________.
29. Calculate the wavelength of a space shuttle with a mass of $2.05 \times 10^7 \text{ kg}$ travelling at a rate of $7.85 \times 10^3 \text{ m/s}$.

Short Answer:

30. What major change in our thinking about the path of a moving electron had to be made with the advent of quantum theory?

---

**Answer Key**

1. c
2. b
3. c
4. b
5. c
6. a
7. d
8. a
9. d
10. b
11. true
12. false
13. false
14. true
15. false
16. false
17. true
18. true
19. false
20. true
21. p
22. five
23. s
24. quantum mechanics
25. wave
26. four
27. six
28. Louis de Broglie
29. \[
\frac{6.626 \times 10^{-34} \text{ J sec}}{2.05 \times 10^7 \text{ kg}(7.85 \times 10^3 \text{ m/sec})} = \frac{\text{kg m}^2/\text{sec}^2}{\text{J sec}} = 4.1 \times 10^{-45} \text{ m}
\]
30. We once thought that electrons moved in simple circular orbits around the nucleus, Quantum theory now requires us to see the electrons moving in much more complicated paths, most of which do not intuitively make sense.
Lesson Quiz

Multiple Choice:

1. The iodine atom has the electronic configuration [Kr]4d^{10}5s^25p^5. The number of valence electrons in this atom is
   a. ten
   b. fifteen
   c. five
   d. seven

2. Lower orbitals are filled before upper orbitals is a statement of the
   a. Hund’s rule
   b. Aufbau principle
   c. Pauli exclusion principle
   d. de Broglie principle

3. The electron configuration for \text{^{12}Mg} is
   a. [Ne]3s^2
   b. [Rn]3s^2
   c. [Ne]3s^23p^2
   d. [Rn]3s^1

4. Electron configuration superscripts indicate
   a. orbital shape
   b. valence electrons
   c. number of electrons in a given sublevel
   d. number of reactive electrons

5. In writing the electron configuration for \text{^{14}Si}, the noble gas to use as a shorthand symbol is
   a. 86Rn
   b. 2He
   c. 10Ne
   d. 18Ar

6. According to the Aufbau principle, the ____ electrons are next in line of filling after the 6s electrons.
   a. 5d
   b. 4f
   c. 4d
   d. 6p

7. Unpaired electrons in orbitals are a consequence of the
   a. Hund’s rule
5.3. Electron Arrangement in Atoms

b. Pauli exclusion principle
c. Aufbau principle
d. Schrödinger equation

8. The first five electron sublevels in order of filling are
   a. 1s, 2s, 3s, 2p, 3p
   b. 1s, 2s, 2p, 3s, 3p
   c. 1s, 2s, 3s, 3p, 2s
   d. 1s, 3s, 2s, 3p, 2p

9. Valence electrons are
   a. unpaired electrons in the outermost principal energy level
   b. all electrons in the outermost orbital
   c. electrons in the outermost principal energy level
   d. electrons with paired spin in the outermost principal energy level

10. The orbital filling for $^3$Li is
    a. ↑↓ ↑
    b. ↑ ↑↓
    c. ↑↑ ↓
    d. ↑↑↓

True/False:

11. _____ An atom’s electron configuration only describes the arrangement of the valence electrons.
12. _____ We can indicate the exact location of any electron.
13. _____ $^3$B will have one unpaired electron.
14. _____ Chlorine ([Ne]3s$^2$3p$^5$) has two unpaired electrons.
15. _____ All unpaired electrons must have the same spin.
16. _____ Orbitals and sub-levels are filled with electrons in order of increasing energy.
17. _____ Neon has one set of unpaired electrons.
18. _____ The electron configuration for $^8$O is 1s$^2$2s$^2$2p$^4$.
19. _____ $^4$Be (1s$^2$2s$^2$2p) has four valence electrons.
20. _____ The sum of the superscripts in an electron configuration is equal to the number of electrons in that atom.

Fill in the Blank:

21. The first ten electrons of the sodium atom are the ________ electrons.
22. In each case, indicate the next sublevel to be filled after the indicated sublevel has all its electrons:
   a. 4s
   b. 5p
   c. 3d
   d. 6s
23. Write the electron configurations for the following atoms:
   a. 19K
   b. 21Sc
   c. 15P
   d. 18Ar
24. Use arrows to indicate the orbital configurations of the following atoms.
   a. selenium: [Ar] 3d$^{10}$4s$^2$4p$^4$ – include only the valence electrons
   b. titanium: [Ar] 3d$^2$4s$^2$
c. niobium: [Kr]4d\textsuperscript{4}5s

25. Two of the isotopes of tungsten are W-180 and W-186. How does the change in the number of neutrons affect the electron configuration of tungsten?

---

Answer Key

1. d
2. b
3. a
4. c
5. c
6. b
7. a
8. b
9. c
10. a
11. false
12. false
13. true
14. true
15. true
16. true
17. false
18. true
19. false
20. true
21. inner-shell
   a. 3d
   b. 6s
   c. 4p
   d. 4f

   a. [Ar]4s
   b. [Ar]3d4s\textsuperscript{2}
   c. [Ne]3s\textsuperscript{2}3p\textsuperscript{5}
   d. [Ne]3s\textsuperscript{2}3p\textsuperscript{6}

\[
\begin{array}{ccc}
\uparrow \downarrow & \uparrow \downarrow & \uparrow \\
4s & 4p \\
\end{array}
\]

\[
\begin{array}{ccc}
\uparrow & \uparrow & \uparrow \\
3d & 4s \\
\end{array}
\]

\[
\begin{array}{ccc}
\uparrow & \uparrow & \uparrow & \uparrow \\
4d & 5s \\
\end{array}
\]

22. The number of neutrons in the nucleus has no effect on the number of electrons or their configuration. The number of protons in the nucleus determines the number of electrons. Hund’s rule, the Aufbau principle, and
the Pauli exclusion principle govern the behavior of the electrons.
Chapter Test

Name___________________ Class______________ Date________

Multiple Choice:

1. Wavelength is defined as
   a. the height of the wave
   b. the distance between two peaks
   c. the frequency of the wave
   d. the speed of the wave

2. One of the following statements about the photoelectric effect is not true
   a. the idea was developed by Albert Einstein
   b. the effect can be seen at all frequencies of light
   c. light can behave as a particle
   d. electrons can be displaced when light shines on a metal surface

3. The spin quantum number gives
   a. the angular momentum of the electron
   b. the rate of rotation of the electron
   c. the direction of spin of the electron
   d. the position of the orbital

4. Chromium has the following electron distribution: [Ar]3d^5\textit{4}s. How many valence electrons does chromium have?
   a. six
   b. one
   c. five
   d. two

5. Wavelength is designated by the symbol
   a. \(\nu\)
   b. \(\delta\)
   c. \(\eta\)
   d. \(\lambda\)

6. Frequency is measured as
   a. cycles/second
   b. cycles/minute
   c. seconds/cycle
   d. vibrations/minute

7. There are _____ possible orientations for p orbitals.
   a. one
5.4. Electrons in Atoms

8. The total number of allowable orbitals in the principal energy level \( n = 3 \) is
   a. 6
   b. 12
   c. 9
   d. 3

9. The units of velocity in the de Broglie wave equation are
   a. km/sec
   b. m/sec
   c. cm/sec
   d. m/minute

10. In writing the electron configuration for \( {}_{24}^{52}\text{Cr} \), the noble gas to use as a shorthand symbol is
   a. \( {}_{36}^{56}\text{Kr} \)
   b. \( {}_{10}^{18}\text{Ne} \)
   c. \( {}_{18}^{54}\text{Ar} \)
   d. \( {}_{54}^{54}\text{Xe} \)

True/False:

11. _____ Unpaired electrons may have different spins.
12. _____ \( {}_{5}^{5}\text{B} (1s^22s^22p) \) has four valence electrons.
13. _____ The p orbital has two lobes.
14. _____ The allowable spin quantum numbers are +2/3 and -2/3.
15. _____ A photon has a greater mass than an electron.
16. _____ The second energy level is filled at the end of the second period.
17. _____ There is no 6f atomic sub-level.
18. _____ \( {}_{18}^{18}\text{Ar} \) has one unpaired electron.
19. _____ A filled orbital has an equal number of electrons spinning in each direction.
20. _____ The Aufbau principle gives the order of electron filling in atoms.

Fill in the Blank:

21. What is the atomic number for the element with the following filling pattern: \( 1s^22s^22p^5 \) _____
22. What is the wavelength of light which has a frequency of 1015 Hz?
23. Write the electron configurations for the following elements:
   a. \( {}_{22}^{52}\text{Ti} \)
   b. \( {}_{19}^{40}\text{K} \)
   c. \( {}_{13}^{27}\text{Al} \)
24. Use arrows to indicate the orbitals for the following –ignore the noble gas component
   a. vanadium: \([\text{Ar}]3d^34s^2\)
   b. germanium (valence electrons only): \([\text{Ar}]3d^{10}4s^24p^2\)
   c. polonium: \([\text{Hg}]6p^4\)
25. Calculate the wavelength of a bumblebee with a mass of \( 1 \times 10^{-3} \) kg travelling at a rate of 2 m/s.

Short Answer:

26. Why would Einstein propose that the electron had properties of a particle?
Answer Key

1. b
2. b
3. c
4. b
5. d
6. a
7. b
8. c
9. b
10. c
11. false
12. false
13. true
14. false
15. false
16. true
17. true
18. false
19. true
20. true
21. nine
22. 

\[ c = \lambda \text{ or } \lambda = \frac{c}{\lambda} \]

\[ = \frac{3 \times 10^8 \text{ m/sec}}{10^{15} \text{ sec}} \]

\[ = 3 \times 10^{-7} \text{ m} \]

a. [Ar]3d\(^2\)4s\(^2\)

b. [Ar]4s

c. [Ne]3s\(^2\)3p

\[ \uparrow \uparrow \uparrow \uparrow \uparrow \]

a. 3d  4s

\[ \uparrow \uparrow \uparrow \uparrow \]

b. 4s  4p

c. 

23. \( \frac{(6.66 \times 10^{-34} \text{ J sec})}{(1 \times 10^{-8} \text{ kg})(2 \text{ m/sec})} \) = \(3.13 \times 10^{-31} \text{ m}\)

24. In order to displace electrons from a metal, the energy striking the metal needed to have mass so that the electrons would be dislodged. There was not enough energy in a massless beam of light at room temperature to move the electrons.
### Chapter Outline

<table>
<thead>
<tr>
<th>Section</th>
<th>Title</th>
</tr>
</thead>
<tbody>
<tr>
<td>6.1</td>
<td>History of the Periodic Table</td>
</tr>
<tr>
<td>6.2</td>
<td>Electron Configuration and the Periodic Table</td>
</tr>
<tr>
<td>6.3</td>
<td>Periodic Trends</td>
</tr>
<tr>
<td>6.4</td>
<td>The Periodic Table</td>
</tr>
</tbody>
</table>
Lesson Quiz

Name___________________ Class________________ Date ________

Multiple Choice:

1. ____________ developed the idea of triads.
   1. Mendeleev
   2. Dobereiner
   3. Newland
   4. Moseley

2. The Law of Octaves stated that every ______ element had similar properties.
   1. eighth
   2. fourth
   3. second
   4. tenth

3. In Mendeleev’s table, elements were arranged __________ in order of increasing atomic mass.
   1. left to right
   2. right to left
   3. top to bottom
   4. bottom top

4. The first element discovered based on Mendeleev’s predictions was
   1. eka-gallium
   2. aluminum
   3. eka-aluminum
   4. gallium

5. The physicist __________ proposed that the periodic table be based on atomic number.
   1. Rutherford
   2. Moseley
   3. Thomson
   4. Newland

6. Each horizontal row of today’s periodic table corresponds to the beginning of a new
   1. principal energy level
   2. spin quantum number
   3. orbital filling
   4. orbital sublevel

7. The vertical columns of the periodic table represent
   1. elements with similar atomic masses
   2. elements with similar atomic radii
3. elements with similar chemical reactivity
4. elements with similar numbers of neutrons

8. The only metal that is a liquid at room temperature is
   1. sodium
   2. radon
   3. mercury
   4. osmium

9. The triad system of classification was based on
   1. atomic number divided by three
   2. both physical and chemical properties
   3. similarities in orbitals
   4. atomic mass multiplied by three

10. The official numbering system for groups uses the numbers
    1. 1-16
    2. 1-20
    3. 1-19
    4. 1-18

True/False:

11. ______ The lanthanide series is listed separately because those atoms are much larger than the others in that period.
12. ______ The Law of Octaves was widely accepted when it was first proposed.
13. ______ When the atomic masses of lithium and potassium are averaged, the resulting number is very close to the atomic mass of sodium.
14. ______ Mendeleev wrote the information about each element on separate note cards.
15. ______ Study of X-ray spectra led to the current definition of atomic number.
16. ______ Elements with similar chemical properties appear in the same horizontal group.
17. ______ The International Chemistry Union developed the current numbering system for the periodic table.
18. ______ Silicon is a typical metalloid.
19. ______ Period two has 18 elements in it.
20. ______ There are four broad classes of elements based on physical properties.

Fill in the Blank:

21. The vertical columns of the periodic table are called ______.
22. The Law of Octaves did not seem to work for elements heavier than ________.
23. Element 101 is named ____________ in honor of the founder of the periodic table.
24. In the modern periodic table, elements are arranged in order of increasing ______ ________.
25. A _____ is a good conductor of heat and electricity.
26. A period is a ___________ row of the periodic table.
27. Metals are ______ and ________.
28. Some new elements have been found in nature, while others have been ________ in the lab.
29. Approximately __________% of the elements in the periodic table are metals.
30. The majority of nonmetals are _____.

Short Answers:

31. Would a computer have helped Mendeleev develop his periodic table? Explain your answer. Assume there was no internet available at the time.
32. A solid has been turned in to the research lab with a request to classify it as a metal, nonmetal, or metalloid. Describe how this could be done.

---

**Answer Key**

1. b
2. a
3. c
4. d
5. b
6. a
7. c
8. c
9. b
10. d
11. false
12. false
13. true
14. true
15. true
16. false
17. false
18. true
19. false
20. false
21. groups
22. calcium
23. mendelevium
24. atomic number
25. metal
26. horizontal
27. malleable, ductile
28. synthesized
29. 80
30. gases
31. Using a spreadsheet or database program, Mendeleev could have quickly tried different ways to organize his data.
32. The following tests and their outcomes are listed

---

**Table 6.1: Class Period(s) (60 min)**

<table>
<thead>
<tr>
<th>Test</th>
<th>Metal</th>
<th>Nonmetal</th>
<th>Metalloid</th>
</tr>
</thead>
<tbody>
<tr>
<td>heat conductivity</td>
<td>good</td>
<td>very poor</td>
<td>some</td>
</tr>
<tr>
<td>electrical conductivity</td>
<td>good</td>
<td>very poor</td>
<td>some</td>
</tr>
<tr>
<td>malleability</td>
<td>very malleable</td>
<td>little malleability</td>
<td>some</td>
</tr>
<tr>
<td>ductility</td>
<td>very ductile</td>
<td>little ductility</td>
<td>some</td>
</tr>
</tbody>
</table>
Lesson Quiz

Name___________________ Class________________ Date ________

Multiple Choice:

1. The first 3s sublevel appears in period
   1. one
   2. two
   3. three
   4. four

2. The ____ sublevel first appears in period five.
   1. 3d
   2. 4s
   3. 4p
   4. 4d

3. The element with the electron configuration [Kr]4d⁵5s² appears in period
   1. seven
   2. five
   3. three
   4. four

4. An element with the electron configuration [Ar]3d⁵4s² is in the _______
   1. f block
   2. d block
   3. p block
   4. s block

5. All of the following are f-block elements except
   1. Re
   2. Sm
   3. Ho
   4. Ce

6. Only one of the following is a p-block element
   1. Te
   2. Ta
   3. Ti
   4. Tm

7. One of the following is not a group one element
   1. Rb
   2. Ca
3. Fr
4. Li

8. The p-block consists of elements in groups
   1. 1-2
   2. 3-12
   3. 13-16
   4. 13-18

9. Group 17 elements are called
   1. inert gases
   2. noble gases
   3. halogens
   4. halides

10. The transition elements have the _____ sublevel being filled
    1. s
    2. p
    3. d
    4. f

True/False:

11. ______ Period six contains 32 elements.
12. ______ The 3d sublevel fills during the fourth period.
13. ______ The s-block elements are largely unreactive.
14. ______ Radium is a group 2 element.
15. ______ Helium has a 1s² electron configuration and is placed in group 18.
16. ______ The p sublevel always fills after the s sublevel of a given principal energy level.
17. ______ In scandium [Ar]3d¹⁴4s², the 3d electron is added before the 4s electron.
18. ______ There are ten elements in each period of the d-block.
19. ______ The actinides consists of nine elements.
20. ______ Many transition element compounds are brightly colored.

Fill in the Blank:

21. Reactions between alkali metals and _____ are extremely vigorous.
22. There are ______ periods on the periodic table.
23. The f sublevel has _____ orbitals.
24. The _____ _____ are in group 18.
25. Alkaline earth elements all have _____ valence electrons, found in the outermost ___ sublevel.
26. Transition elements found pure in nature are ______, ______, and ______.
27. The halogens react vigorously with the _____ metals.
28. Compounds of xenon have been formed through reaction with _______ gas.
29. Fill in the blanks

<table>
<thead>
<tr>
<th>Configuration</th>
<th>Period</th>
<th>Group</th>
<th>Block</th>
<th>Valence Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Xe⁴f¹⁴⁵d¹⁶s²</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>[Ne]³s²³p</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>1s²2s²2p³</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Short Answer:

30. Why is helium a group 18 element?

---

**Answer Key**

1. c
2. d
3. b
4. b
5. a
6. a
7. b
8. d
9. c
10. c
11. true
12. true
13. false
14. true
15. true
16. true
17. false
18. true
19. false
20. true
21. halogens
22. seven
23. seven
24. noble gases
25. two, s
26. platinum, gold, silver
27. alkali
28. fluorine

---

**Table 6.3:** Class Period(s) (60 min)

<table>
<thead>
<tr>
<th>Configuration</th>
<th>Period</th>
<th>Group</th>
<th>Block</th>
<th>Valence Electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>[Xe]4f¹⁴5d⁵6s²</td>
<td>six</td>
<td>six</td>
<td>d</td>
<td>two</td>
</tr>
<tr>
<td>[Ne]3s²3p</td>
<td>three</td>
<td>thirteen</td>
<td>p</td>
<td>three</td>
</tr>
<tr>
<td>1s²2s²2p³</td>
<td>two</td>
<td>fifteen</td>
<td>p</td>
<td>five</td>
</tr>
</tbody>
</table>

30. Although He only has s electrons, its 1s² configuration gives it a complete outer orbital. Chemically, He behaves like the noble gases since it is very unreactive.
Lesson Quiz

Multiple Choice:

1. The atomic radius is measured in
   1. picometers
   2. nanometers
   3. femtometers
   4. micrometers

2. The outer electron configuration for group 14 is
   1. ns²
   2. np⁴
   3. ns²p³
   4. ns²p⁴

3. When oxygen forms an ion, the name ends in
   1. ate
   2. ite
   3. ode
   4. ide

4. Ionization energy is the energy needed to
   1. remove an electron from an atom
   2. add an electron to an atom
   3. add an electron to an ion
   4. remove a proton from an atom

5. For each period in the periodic table, the ionization energy is highest in group
   1. 15
   2. 18
   3. 13
   4. 10

6. The elements of group ____ gain electrons most readily.
   1. 4
   2. 7
   3. 13
   4. 17

7. Electron shielding is
   1. blocking inner orbital electron removal by outer electrons
   2. inner electron partially shielding outer electrons from proton charge
   3. blocking of protons from inner electron attractions
   4. p orbital blocking of s orbital removal
8. Electron affinity is the
   1. energy released when an electron is lost
   2. energy released when electrons change orbitals
   3. energy released when an atom gains an electron
   4. energy needed to release an electron from an atom

9. The atomic radius
   1. decreases from top to bottom on the periodic table
   2. increases from left to right across the periodic table
   3. increases from top to bottom on the periodic table
   4. is greatest in the middle of the periodic table

10. All of the following statements about electronegativity are true with one exception
    1. electronegativity is not a measured unit of energy
    2. involves repulsion of electrons on a compound
    3. fluorine is the most electronegative atom
    4. electronegativity values for metals are generally low

True/False:

11. ______ For group one elements, the ionic radius increases as the atomic number increases.
12. ______ Elements within a group share similar chemical properties.
13. ______ Some valence electrons are in d and f orbitals.
14. ______ Ions form by atoms losing or gaining electrons.
15. ______ The sulfide ion is formed when sulfur loses two electrons.
16. ______ It is easier to remove a paired electron than an unpaired electron.
17. ______ The second ionization energy for an element is always less than the first ionization energy.
18. ______ The radius of the calcium ion is less than the radius of the calcium atom.
19. ______ The noble gases have very high electronegativity values.
20. ______ The electron affinities of group 17 elements is much higher than those of group two elements.

Fill in the Blank:

21. A positively charged ion is called a ______ and a negatively charge ion is called an ______.
22. As the atomic _____ increases within a period, the atomic ______ decreases.
23. The outer electron configuration for group 16 is ________.
24. The chemical symbol for the magnesium ion is ______.
25. Ionization energies are measured in _______.
26. Ionization energies generally ________ from top to bottom within a group.
27. The removal of electrons always results in a cation that is ______ than the parent atom.
28. ______ attracts electrons better than any other element.
29. The distance between two adjacent aluminum atoms is 286 pm. The atomic radius of the AL atom is ______- __ pm.

Short Answers:

30. The potassium atom has 19 protons and 19 electrons. What will potassium most likely do to form an ion? The following patterns are observed with regard to ionization energies (IE):
### Table 6.4: Class Period(s) (60 min)

<table>
<thead>
<tr>
<th>Element</th>
<th>IE1</th>
<th>IE2</th>
<th>IE3</th>
<th>IE4</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>520</td>
<td>7300</td>
<td>——</td>
<td>——</td>
</tr>
<tr>
<td>B</td>
<td>801</td>
<td>2430</td>
<td>3660</td>
<td>25,000</td>
</tr>
<tr>
<td>Mg</td>
<td>738</td>
<td>1450</td>
<td>7730</td>
<td>——</td>
</tr>
</tbody>
</table>

Explain why the ionization energy suddenly increases significantly in each case.

**Answers Key**

1. a
2. d
3. d
4. a
5. b
6. d
7. b
8. c
9. c
10. b
11. true
12. true
13. false
14. true
15. false
16. true
17. false
18. true
19. false
20. true
21. cation, anion
22. number, radius
23. ns²np⁴
24. Mg²⁺
25. kJ/mol
26. decrease
27. smaller
28. fluorine
29. 143
30. lose an electron
31. In each case, the lower ionization energies are seen as the valence electrons are removed. For example, Li has one valence electron (1s²2s). When the 2s valence electron is removed, the resulting configuration is now similar to that of a noble gas.
6.4 The Periodic Table

Chapter Test

Name___________________ Class______________ Date ________

Multiple Choice:

1. Ions are formed by
   1. only loss of electrons
   2. loss or gain of electrons
   3. only gain of electrons
   4. simultaneous loss and gain of electrons

2. $^{40}\text{Zr}$ belongs in the
   1. s-group
   2. p-group
   3. d-group
   4. f-group

3. Mendeleev organized his periodic table based on
   1. electron orbitals
   2. electron number
   3. atomic number
   4. atomic mass

4. Ra is in the _______ group.
   1. alkali metals
   2. alkaline earth metals
   3. halogens
   4. noble gases

5. A group in the periodic table is
   1. a horizontal row of elements
   2. adjacent periods of elements
   3. a vertical column of elements
   4. adjacent columns of elements

6. The d sublevel has ____ electrons
   1. 6
   2. 2
   3. 10
   4. 14

7. The lanthanides are found in group _____ of the periodic table
   1. 5
   2. 3
3. 7
4. 10

8. Moseley made the following improvement to Mendeleev’s periodic table
   1. used atomic numbers instead of atomic mass
   2. developed more accurate atomic mass values
   3. based the table on electron configurations
   4. organized the table based on protons plus neutrons

9. The group 14 elements have the electron configuration
   1. ns²np³
   2. nsnp⁴
   3. ns²np²
   4. nd¹⁰s²p²

10. Period 5 of the periodic table contains ______ elements
    1. 12
    2. 6
    3. 24
    4. 18

True/False:

11. ______ Ionization energy increases markedly after all valence electrons have been removed.
12. ______ Metalloids have one less valence electrons than metals in the same period.
13. ______ Newland’s idea of octaves did not work well for elements with atomic numbers higher than calcium.
14. ______ ⁷⁰Yb is a p-block element
15. ______ The ion formed when Be loses an electron has a smaller atomic radius than the parent atom.
16. ______ As the atomic number increases within a group, the atomic radius also increases.
17. ______ A nonmetal has a high electrical conductivity.
18. ______ The number of valence electrons and the outer electron configuration are constant within a group.
19. ______ All noble gases have the ns²np⁶ electron configuration.
20. ______ Chlorine is the most electronegative element.

Fill in the Blank:

Use a periodic table to help answer the following questions

21. All group 12 elements have a ______ valence electron configuration.
22. The period 7 group 3 elements are known as ______ and are characterized by having unfilled _____ and _____ orbitals.
23. More energy is released in the formation of a ______ ion than for the anions of any other elements.
24. Dobereiner’s proposal about triads was the precursor of the idea of similar _____ of elements within groups.
25. ______ Which halogen has an atomic number of 53 and an atomic mass of 126.0? ______
26. The inner transition elements are known as the ______ and _________. The majority of these elements have unfilled _____ sublevels.
27. ______ The electron configuration [Xe]⁴f¹⁴⁵d⁹⁶s represents the element _____.
28. ______ Write the electron configuration for ¹₅₃P

Short Answers:

29. ______ Explain why all valence electrons are either in an s or p orbital.
30. ______ The first ionization energy for ⁵B is 801 kJ/mol while the first ionization energy for ¹₃Al is 578 kJ/mol. Offer a reasonable explanation for this observation.
Answers Key

1. b
2. c
3. d
4. b
5. c
6. c
7. b
8. a
9. c
10. d
11. true
12. false
13. true
14. false
15. false
16. false
17. false
18. true
19. true
20. false
21. s²
22. actinides, d, f
23. halide
24. reactivities
25. iodine
26. lanthanides, actinides, f
27. platinum or Pt
28. [Ne]3s²3p³
29. d and f orbitals fill after the higher s and p orbitals.
30. The first electron to be removed for Al is somewhat further away from the nucleus than in the case of B. In addition, inner s and p electrons in aluminum provide electron shielding, making that first valence electron for Al easier to remove.
Chapter 7

Chemical Nomenclature Assessments

Chapter Outline

7.1 Ionic Compounds
7.2 Molecular Compounds
7.3 Acids and Bases
7.4 Chemical Nomenclature
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. A molecular formula shows
   1. how many atoms are in the molecule
   2. the arrangement of atoms in a molecule
   3. the lowest whole-number ratio of atoms in a molecule
   4. the three-dimension array of ions in a molecule

2. One of the following is not a monatomic ion
   1. Na⁺
   2. Cl⁻
   3. ClO⁻
   4. O²⁻

3. The oxide ion has a charge of
   1. +1
   2. +2
   3. -1
   4. -2

4. The magnesium ion has a charge of
   1. +1
   2. +2
   3. -1
   4. -2

5. The Stock system is used in naming compounds containing
   1. metals
   2. cations
   3. transition metals
   4. transition anions

6. The correct name for the As³⁻ ion is
   1. arsenic
   2. arsenous
   3. arsenide
   4. arsenate

7. The correct charge for the chromium(III) ion is
   1. Cr⁺
   2. Cr²⁺
3. Cr$^{3+}$
4. Cr$^{4+}$

8. _____ is the name for the HPO$_4^{2-}$ ion
   1. monohydrogen phosphate
   2. hydrogen phosphite
   3. biphosphate
   4. hydrogen phosphate

9. Parentheses are used in ionic formulas when
   1. ionic symbols equal one another
   2. more than one polyatomic ion is present
   3. separating cations and anions
   4. when more than one monatomic ion is present

10. Hg$^{2+}$ is a
    1. polyatomic cation
    2. polyatomic compound
    3. monatomic cation
    4. trivalent cation

True/False:

11. _____ Na$_2$Cl$_2$ is the correct empirical formula for sodium chloride.
12. _____ Ionic compounds exist in three-dimensional arrays of ions.
13. _____ The proper designation of charge for the potassium ion is K$^{+1}$.
14. _____ HCO$_3^-$ is the symbol for the bicarbonate anion.
15. _____ The peroxide anion is monatomic.
16. _____ Chlorine can form four different anions with oxygen.
17. _____ The final formula for a ternary ionic compound must be neutral.
18. _____ Zn$_6$(PO$_4$)$_4$ is the correct empirical formula for zinc phosphate.
19. _____ All carbon-containing compounds are considered to be organic.
20. _____ A ternary compound consists of three or more elements.

Fill in the Blank:

21. Name the following compounds:
   1. Ag$_2$S
   2. Fe$_2$O$_3$
   3. CuCl$_2$
   4. AuBr
   5. MnO
   6. Ag$_3$SO$_4$

22. Write formulas for the following compounds
   1. lead(II) nitride
   2. cobalt(III) selenide
   3. rubidium dichromate
   4. magnesium cyanide
   5. lithium hypochlorite
   6. ammonium arsenate

Short Answer:
23. What is the advantage of using a chemical formula instead of a chemical name?

**Answer Key**

1. a
2. c
3. d
4. b
5. c
6. c
7. c
8. d
9. b
10. a
11. false
12. true
13. false
14. true
15. false
16. true
17. true
18. false
19. false
20. true

1. silver sulfide
2. iron(III) oxide
3. copper(II) chloride
4. gold(I) bromide
5. manganese(II) oxide
6. silver sulfate

1. Pb₃N₂
2. Co₂Se₃
3. Rb₂Cr₂O₇
4. Mg(CN)₂
5. LiClO
6. (NH₄)₃AsO₃

21. The chemical symbols show relationships among atoms. These symbols are a universal language, understood by all people. Seeing relationships among compounds is much easier using symbols. Calculations are easier.
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. Only one of the following statements about molecular compounds is true:
   1. molecular compounds exist in extended arrays
   2. molecular compounds are made up of alternating positive and negative ions
   3. molecular compounds form bonds between pairs of atoms
   4. molecular compounds are represented by empirical formulas

2. Which of the following is a binary molecular compound?
   1. NaCl
   2. CH₃Cl
   3. Na₂Cr₂O₇
   4. CH₃CH₂CH₃

3. The prefix ______ could designate a compound with six carbon atoms.
   1. tetra
   2. hexa
   3. octa
   4. deca

4. The compound tribromopentane contains ______ bromine atoms
   1. two
   2. three
   3. four
   4. five

5. The element ___ follows Cl when writing binary compound formulas.
   1. O
   2. H
   3. Br
   4. S

6. The formula for sulfur dioxide is
   1. SO
   2. S₂O
   3. SO₂
   4. S₂O₂

7. The formula P₂O₅ represents the compound
   1. pentoxide diphosphate
   2. diphosphorus pentoxide
3. phosphorus pentaoxide
4. diposphopentoxide

8. The formula BH₃ represents the compound
   1. boron trihydride
   2. beryllium trihydride
   3. boron trihydride
   4. boron trihydrogen

9. Tellurium trioxide is represented by the following formula
   1. Te₂O₃
   2. TeO₃
   3. Te₃O₂
   4. TeO₄

10. Oxygen difluoride has the formula
   1. OF₃
   2. O₃F
   3. O₃F₃
   4. OF₂

True/False:

11. _____ Carbon tetrachloride is an ionic compound.
12. _____ Carbon monoxide contains one oxygen atom.
13. _____ Cl and O can combine in a variety of ratios.
14. _____ Generally, the most electronegative element is written first in a formula.
15. _____ The prefix mono is always used if there is one atom of the first element.
16. _____ The ending of the second element’s name is changed to -ide.
17. _____ Four oxygen atoms is always referred to as tetraoxide.
18. _____ Metalloids generally form molecular compounds.
19. _____ Hepta- is the prefix used to designate six atoms of a specific element in a compound.
20. _____ The name dinitrogen monoxide suggests a compound containing two nitrogen atoms and one oxygen atom.

21. Name the following compounds:
   1. S₂Cl₂
   2. NO₂
   3. CCl₄
   4. IBr₄
   5. BrF₅
   6. SiO₃

22. Write formulas for the following compounds:
   1. bromochloride
   2. oxygen difluoride
   3. phosphorus trihydride
   4. diiodine pentoxide
   5. tricarbon octabromide
   6. disulfur difluoride

Short Answer:

23. How are the atoms held together in molecular compounds? How does this differ from attraction in ionic compounds?
1. disulfur dichloride
2. nitrogen dioxide
3. carbon tetrachloride
4. iodine tetrabromide
5. bromine pentafluoride
6. silicon trioxide
7. BrCl
8. OF₂
9. PH₃
10. I₂O₅
11. C₃Br₈
12. S₂F₂

21. Atoms in molecular compound share electrons and are held together by those attractions. Ionic compounds have charge attractions that hold the ions together in large clusters or arrays.
7.3 Acids and Bases

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. ______ produce H\(^+\) when dissolved in water.
   1. acids
   2. molecular compounds
   3. bases
   4. ionic compounds

2. A ______ produces hydroxide ions when dissolved in water.
   1. acid
   2. molecular compounds
   3. base
   4. ionic compounds

3. A base contains at least one ____ atom
   1. Na
   2. O
   3. K
   4. Cl

4. One of the following does not contain acids
   1. citrus fruits
   2. citrus juices
   3. drain cleaners
   4. vinegar

5. Acid structure is
   1. extended
   2. three-dimensional
   3. ionic
   4. molecular

6. Base structure is
   1. molecular
   2. covalent
   3. ionic
   4. shared electrons

7. NH\(_4\)OH is properly named
   1. nitrogen tetrahydride hydroxide
   2. ammonium hydroxide
3. ammonium basic oxide
4. nitrogen hydroxide

8. The proper name for HCl in water is
   1. hydrogen chloride
   2. hydrochloride
   3. hydrochloric acid
   4. hydride chloric acid

9. When the anion ends in -ite, the suffix for the acid is
   1. -us
   2. -ic
   3. -ate
   4. -ous

10. It is the _____ of the anion that determines how the acid is named.
    1. suffix
    2. prefix
    3. beginning
    4. middle

True/False:

11. _____ A binary acid is composed of hydrogen and one other element.
12. _____ CH₃Cl is an acid
13. _____ All monatomic anions end in -ide.
14. _____ The acid prefix for an anion ending in -ide is hypo-.
15. _____ The proper name for Ca(OH)₂ is calcium dihydroxide.
16. _____ The hydroxide ion is polyatomic.
17. _____ Acids are electrically neutral.
18. _____ H₂SO₄ is called hydrosulfic acid.
19. _____ All inorganic bases are called hydroxides.
20. _____ An -ite anion requires the use of -ous in naming the acid.

Fill in the Blank:

21. Name the following compounds:
   1. HBr
   2. HNO₂
   3. Ba(OH)₂
   4. RbOH
   5. H₃PO₃
   6. H₂CrO₄

22. Write the formulas for the following compounds:
   1. carbonic acid
   2. hydroiodic acid
   3. strontium hydroxide
   4. iron(II) hydroxide
   5. permanganic acid
   6. chloric acid

Short Answer:
23. Why does the definition of an acid include two parts? Is it enough to say that acids contain hydrogen atoms?

**Answer Key**

1. a  
2. c  
3. b  
4. c  
5. d  
6. c  
7. b  
8. c  
9. d  
10. a  
11. true  
12. false  
13. true  
14. false  
15. false  
16. true  
17. true  
18. false  
19. true  
20. true

1. hydrobromic acid  
2. nitrous acid  
3. barium hydroxide  
4. rubidium hydroxide  
5. phosphorus acid  
6. chromic acid

1. $\text{H}_2\text{CO}_3$  
2. HI  
3. $\text{Sr(OH)}_2$  
4. $\text{Fe(OH)}_2$  
5. $\text{HMnO}_4$  
6. $\text{HClO}_3$

21. An acid must contain H atoms and it must form hydrogen ions in solution. Many compounds contain hydrogen atoms, but do not form hydrogen ion when dissolved in water.
Chapter Test

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. An acid contains at least one ____ atom.
   1. O
   2. Cl
   3. H
   4. N

2. The Stock system is used with
   1. transition metals
   2. noble gases
   3. alkali metals
   4. halogens

3. The molecular formula shows
   1. the shape of the molecule
   2. the number of atoms in the molecule
   3. the number of ions in the molecule
   4. the whole number ratio of atoms in the molecule

4. Cations are generally formed by
   1. removing some of the valence electrons
   2. removing only one of the valence electrons
   3. removing all the valence electrons
   4. adding to the valence electron sublevel

5. Transition metals
   1. are found in groups 1 and 2
   2. form more than one type of stable cation
   3. are all f block elements
   4. are atomically unstable

6. Oxygen trifluoride has the following formula
   1. OF$_3$
   2. O$_3$F
   3. O$_3$F$_3$
   4. OF$_2$

7. Selenium dioxide has the following formula
   1. SeO$_3$
   2. Se$_3$O$_2$
3. SeO$_2$
4. Se$_2$O$_2$

8. When dissolved in water, HCl gas will
1. stay unchanged
2. form extended arrays of HCl molecules
3. form HCl$^+$ ions
4. form H$^+$ and Cl$^-$

9. A binary compound
1. contains two types of atoms
2. contains atoms combined in ratios of two
3. is a compound made of two metals
4. is a compound made of two nonmetals

10. The correct name for Sn$_3$P$_2$ is
1. tin phosphide
2. tin(II) phosphate
3. tin(II) phosphide
4. tin phosphate

True/False:

11. _____ KMnO$_4$ is a ternary compound.
12. _____ Ca(II) is the correct way to name the Ca$^{2+}$ ion.
13. _____ All binary compounds are electrically neutral.
14. _____ A polyatomic ion has a charge of +2 or greater.
15. _____ Carbon dioxide has two carbons in between two oxygens.
16. _____ The formula for carbon tetrabromide is CBr$_4$.
17. _____ A base produces hydronium ions when dissolved in water.
18. _____ All monatomic ions end in ate.
19. _____ CN$^-$ is the cyanate anion.
20. _____ AgCl is known as silver chloride.

Fill in the Blank:

21. Name the following compounds
1. NaClO$_3$
2. K$_2$SO$_3$
3. H$_3$PO$_4$
4. MnCl$_2$
5. AuI$_3$
6. Ba(NO$_3$)$_2$
7. H$_2$SO$_4$

22. Write formulas for the following compounds
1. dinitrogen pentoxide
2. nickel(II) chloride
3. iron(III) silicate
4. hydroiodic acid
5. chromium(II) hydroxide
6. carbon tetranitrate
7. ammonium cyanide

Short Answer:

23. Why is it so important to have the exact name and formula for a compound?

Answer Key

1. c
2. a
3. b
4. c
5. b
6. a
7. c
8. d
9. a
10. c
11. true
12. false
13. true
14. false
15. false
16. true
17. false
18. false
19. false
20. true

1. sodium chlorate
2. potassium sulfite
3. phosphoric acid
4. manganese(II) chloride
5. gold(III) iodide
6. barium nitrate
7. sulfuric acid
8. N₂O₅
9. NiCl₂
10. Fe₂(SiO₃)₃
11. HI
12. Cr(OH)₂
13. C(NO₃)₄
14. NH₄CN

21. Reactions involve specific materials in order to occur. The wrong material can produce misunderstanding when trying to demonstrate chemical principles. Adverse reactions can occur, including hazardous situations. In medicine, giving the wrong drug to a patient will not help their recovery and can seriously harm them.
Chapter 8

Ionic and Metallic Bonding

Assessments

Chapter Outline

8.1 IONS
8.2 IONIC BONDS AND IONIC COMPOUNDS
8.3 METALLIC BONDS
8.4 IONIC AND METALLIC BONDING
8.1 Ions

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. All of the elements below have two valence electrons except
   1. Sr
   2. Ra
   3. Cr
   4. Be

2. Tc has ______ valence electrons
   1. 2
   2. 5
   3. 7
   4. 9

3. Which of the following transition metal ions is not particularly stable?
   1. Co$^{2+}$
   2. Cu$^{+}$
   3. Ag$^{+}$
   4. Zn$^{2+}$

4. Anions are formed when
   1. atoms lose one electron
   2. atoms gain one electron
   3. atoms lose their valence shell electrons
   4. atoms add electrons to fill their valence shell

5. In order to achieve a full octet, the arsenic atom will most likely
   1. lose 3 electrons
   2. gain 3 electrons
   3. lose 5 electrons
   4. lose 2 electrons

6. The electron configuration for the Mo$^{+}$ ion is
   1. [Kr]4d$^{5}$5s
   2. [Xe]4d$^{5}$
   3. [Kr]4d$^{4}$
   4. [Kr]4d$^{5}$

7. Atoms of ______ tend to gain electrons to satisfy the octet rule.
   1. gases
   2. halogens
3. metals
4. noble gases

8. The $P^{3-}$ anion is isoelectric with
   1. Ne
   2. Xe
   3. Kr
   4. Ar

9. S-block elements have a maximum of ___ valence electrons
   1. 1
   2. 3
   3. 2
   4. 4

10. is the electron dot structure for which set of elements?
   1. halogen
   2. group one
   3. noble gases
   4. group two

True/False:

11. _____ [Ar]$3d^34s^2$ is the electron configuration for vanadium.
12. _____ Mg$^{2+}$ and Ar are isoelectronic.
13. _____ The pseudo noble gas electron configuration for some transition elements contains 16 electrons.
14. _____ All the period six elements have a filled 4f principal energy level.
15. _____ Barium can achieve a complete outermost principal energy level by losing two electrons.
16. _____ Linus Pauling first proposed the octet rule.
17. _____ Zr has two valence electrons.
18. _____ The electron dot diagram for Cs will have two dots in it.
19. _____ Electron dot diagrams for a given group are identical (except for the element symbol).
20. _____ The octet rule indicates that elements form compounds in ways that give each element eight electrons.

Fill in the Blank:

21. Which cation is isoelectronic with Cl$^-$?
22. Draw electron dot diagrams for
   1. P
   2. Po
   3. As
23. Write the ion formed by each of the following elements in order to fulfill the octet rule:
   1. B
   2. I
   3. Zn
   4. Be
24. Write the electron configurations for the following ions:
   1. Ag$^+$
   2. Se$^{2-}$
3. Tc$^{2+}$
4. Te$^{2-}$

Short Answer:
25. Do the transition elements always follow the Aufbau rules?

Answer Key

1. c
2. a
3. a
4. d
5. b
6. d
7. b
8. d
9. c
10. c
11. true
12. false
13. false
14. false
15. true
16. false
17. true
18. false
19. true
20. true
21. K$^+$

1. $\ddot{P}$
2. $\ddot{P}$
3. $\ddot{S}$

1. B$^{3+}$
2. I$^-$
3. Zn$^{2+}$
4. Be$^{2+}$

1. [Kr]4d$^{10}$
2. [Ar]3d$^{10}$4s$^2$4p$^6$
3. [Kr]4d$^5$
4. [Kr]4d$^{10}$5s$^2$5p$^6$

22. No, they don’t. Example: V has an electron configuration of [Ar]3d$^3$4s$^2$. One would expect the next element to add another 3d electron since the 4s energy level is full. Instead, Cr has an electron configuration of [Ar]3d$^5$4s. Other examples can easily be found.
Lesson Quiz

Multiple Choice:

1. One of the following statements is not true
   1. the K-Cl ionic bond is weaker than the Na-Cl ionic bond
   2. the Ra-Cl ionic bond is stronger than the Be-Cl ionic bond
   3. the Li-Cl ionic bond is stronger than the Rb-Cl ionic bond
   4. the Mg-Cl ionic bond is stronger than the Ba-Cl ionic bond

2. Oxygen can most easily form an ion by
   1. adding two electrons
   2. adding four electrons
   3. losing two electrons
   4. losing four electrons

3. \(\text{:\text{X}:}\) is the electron dot symbol for the group two element
   1. Be
   2. S
   3. P
   4. Li

4. The formula unit for aluminum sulfate is
   1. AlSO_4
   2. Al(SO_4)_2
   3. Al_2(SO_4)_3
   4. Al_3(SO_4)_2

5. Models of ionic compounds can be illustrated using a _______ model
   1. ball and stick
   2. crystal lattice
   3. ball and chain
   4. stick crystal

6. Only one of the following compounds will not conduct electricity when melted
   1. KCl
   2. CaBr_2
   3. CH_3Br
   4. NaI

7. Which of the following ions will form the weakest bond with Na^+

1. F⁻
2. I⁻
3. Cl⁻
4. Br⁻

8. The nitride anion involves the addition of ____ electrons to the N atom.
   1. 2
   2. 3
   3. 1
   4. 4

9. Ionic compounds exist in extended arrays to
   1. make stacking easier
   2. balance charges
   3. minimize potential energy of the system
   4. optimize electron transfer

10. The most accurate representation of ion packing is the
    1. ball and stick model
    2. space filling array model
    3. space balancing array model
    4. ball and rod model

True/False:

11. _____ Copper can lose two electrons to form an ion.
12. _____ One formula unit of \( \text{Al}_2(\text{SO}_4)_3 \) contains two sulfates for every aluminum.
13. _____ All group 2 cations will have the same packing arrangement.
14. _____ The coordination number tells how many ions surround an ion of opposite charge.
15. _____ Coordination numbers for the cation and anion in a crystal are identical.
16. _____ Transition ions give rise to color in crystals.
17. _____ Melted CsCl can conduct electricity.
18. _____ Ions of the same charge repel one another.
19. _____ Ion crystals form rough jagged edges when they break.
20. _____

Fill in the Blank:

21. LiCl and PbCl₂ are both ionic compounds. Which cation will more strongly attract the chloride ions? Explain your answer.
22. Iodine is a _______ and _____ one electron to become an anion.
23. The formula unit of an ionic compound is always an _______ formula.
24. _____ and ____ alternate in the extended three-dimensional array of an ionic compound.
25. Why do CsCl and LiCl not have identical packing arrangements?
26. What type of packing arrangement does CoCl₂ have when dissolved in water?
27. _______ charged particles attract one another.
28. An _____ compound is an electrically _____ compound consisting of _____ and _____ ions.

: X :

is the period 4 electron dot formula for germanium.
8.2. Ionic Bonds and Ionic Compounds

Answer Key

1. b
2. a
3. b
4. c
5. a
6. c
7. b
8. b
9. c
10. c
11. true
12. false
13. false
14. true
15. false
16. true
17. true
18. true
19. false
20. True
21. Li will because its ionic radius is much smaller than that of lead.
22. halogen, adds
23. empirical
24. Anions, cations
25. The sizes of the two ions are very different. So that will affect the arrangements of anions around the central ion.
26. Oppositely
27. ionic, neutral, anions, cations
**8.3 Metallic Bonds**

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. A metallic bond is
   1. the attraction of metal ions to one another
   2. the attraction of metal ions to mobile electrons
   3. the force generated by stationary electrons
   4. a bond between a metallic cation and an anion

2. Metals conduct heat because
   1. electrons vibrate freely
   2. electrons are easily elevated to higher energy levels
   3. electrons release energy when they drop to lower energy levels
   4. electrons flow freely through the metal

3. Atoms in metals
   1. are stationary
   2. vibrate freely
   3. flow through the metal surface
   4. experience periodic loss of electricity

4. The _________ metal crystal has the greatest amount of space between atoms
   1. hexagonal
   2. polyhedral
   3. cubic face centered
   4. cubic body centered

5. Which two metal crystal structures have the same coordination number?
   1. body-centered cubic and hexagonal
   2. face-centered cubic and hexagonal
   3. body-centered cubic and face-centered cubic
   4. face-centered cubic and polygonal

6. Brass is an alloy composed of
   1. copper and tin
   2. zinc and tin
   3. zinc and vanadium
   4. copper and zinc

7. Only one of the following is a constituent of steel
   1. copper
   2. boron
8.3. Metallic Bonds

3. strontium
4. calcium

8. One of the following is not a property of metals
   1. high ductility
   2. high malleability
   3. high melting point
   4. high compressibility

9. The photoelectric effect contributes to the _____ of metals
   1. luster
   2. malleability
   3. conductivity
   4. ductility

10. The cubic body-centered crystal structure is seen in the metal ______.
    1. Al
    2. Cd
    3. Cr
    4. Cu

True/False:

11. _____ The major constituent of steel is iron.
12. _____ Brass is an alloy of copper and tin.
13. _____ Close packing provides room for the electrons to flow easily.
14. _____ Brass is used in the manufacture of trumpets.
15. _____ Se shows a high degree of metallic bonding.
16. _____ The hexagonal packing arrangement can be seen in crystals of Mg.
17. _____ Metallic materials are very brittle.
18. _____ Metal cations are surrounded by mobile electrons.
19. _____ Metal electrons are primarily associated with specific atoms.
20. _____ In the body-centered cubic structure, one metal atom is in the very center of the structure.

Fill in the Blank:

21. An alloy is a mixture of _____ or more elements, at least one of which is a ______.
22. Name the crystal structures:

23. A
24. A: _________________________________ B: _________________________________ C: ________________

25. The model of metallic bonding is called the ___________ model.

26. Alloys are commonly used in ________ objects because the properties of the mixtures are often more _______ than the pure metal.

27. Write definitions for the following terms:
   1. substitutional alloy
   2. interstitial alloy
   3. bronze
   4. brass
   5. closest packing

Short Answer:

26. Which energy level electrons would you expect to be most mobile? Explain your answer.

Answer Key

1. b
2. d
3. a
4. d
5. b
6. d
7. b
8. d
9. a
10. c
11. true
12. false
13. false
14. true
15. false
16. true
17. false
18. true
19. false
20. true
21. two, metal
   1. cubic body-centered
   2. cubic face centered
   3. hexagonal
22. sea of mobile electrons
23. manufactured, useful
   1. The various atoms simply replace each other in the crystal structure.
   2. Smaller atoms such as carbon fit in between the larger atoms in the crystal packing arrangement.
   3. an alloy of copper and tin.
   4. an alloy of copper and zinc.
   5. the most efficient way to pack spherical objects.
24. The outermost principal energy level should be the most mobile. The electrons ae easier to “detach” from the atom since there are other inner electrons that can serve to shiled the outer electrons from nuclear attraction.
Chapter Test

Name _____________________  Class ______________________  Date __________________

Multiple Choice:

1. A +2 charge will be seen when ___ forms an ion
   1. Na
   2. O
   3. Mg
   4. Rn

2. The electron configuration for _____ is [Ne]3s²3p⁶
   1. Na⁺
   2. S²⁻
   3. Mg²⁺
   4. P²⁻

3. The formula unit for aluminum sulfate is
   1. AlSO₄
   2. Al₃(SO₄)₂
   3. Al(SO₄)₂
   4. Al₂(SO₄)₃

4. The coordination number for an ionic compound is
   1. the number of ions in the crystal
   2. the number of ions in the molecule
   3. the number of ions immediately surrounding an ion of opposite charge
   4. the number of ions immediately surrounding an ion of like charge

5. \( \cdot X \cdot \)
   is the electron dot symbol for
   1. C
   2. Al
   3. Au
   4. Mn

6. The ball and stick model shows
   1. shared electrons
   2. cation relationships
   3. ion packing
   4. ion distribution

7. Which of the following ions will form the weakest ionic bond with Cl?
   1. Cs⁺
8.4. Ionic and Metallic Bonding

2. Ca$^{2+}$
3. Na$^{+}$
4. Be$^{2+}$

8. Metals conduct electricity because
   1. electrons vibrate freely
   2. electrons are easily elevated to higher energy levels
   3. electrons release energy when they drop to lower energy levels
   4. electrons flow freely through the metal

9. Cations are formed by
   1. adding electrons to the valence shell
   2. adding electrons to lower energy levels
   3. losing all their valence electrons
   4. losing lower energy level electrons

10. A charge of ____ is very common for transition metal ions
    1. 2$^-$
    2. 2$^+$
    3. 3$^+$
    4. 3$^-$

**True/False:**

11. _____ The crystal structure of any ionic compound must reflect its formula unit
12. _____ Co can lose two electrons to form an ion
13. _____ A half-filled d sublevel is very unstable.
14. _____ NaCl has a high melting point
15. _____ Melted CaBr will not conduct electricity.
16. _____ The crystal structure of a compound must reflect its formula unit.
17. _____ Cr has one valence electron.
18. _____

   is the electron dot formula for a halogen atom.

19. _____ Under typical conditions, a maximum of three electrons will be gained during anion formation.
20. _____ Fe$_2$O$_3$ is the principal component of rust.

**Fill in the Blank:**

22. Draw the electron dot diagram for
   1. oxygen
   2. Ca
   3. He
   4. Sr

23. Indicate whether electrons will be lost or gained when the following elements form ions and state the number of electrons involved.
   1. H
   2. Hf
3. Po
4. Se

24. A ___________ model of crystal structure shows the ions in contact with each other.

Short Answer:

25. Explain why Fe can form a 3+ cation.

---

**Answer Key**

1. c
2. b
3. d
4. c
5. b
6. d
7. a
8. d
9. c
10. b
11. true
12. true
13. false
14. true
15. false
16. true
17. true
18. false
19. true
20. true
21. 18

22. space-filling

23. The iron atom has the electron configuration [Ar]3d⁶⁴s². A 2+ cation can be formed by losing the two 4s electrons, leaving a 3d⁶ outer energy level. Since a half-filled level would be more stable, the 3+ ion can form by losing a 3d electron and forming the stable 3d⁵ configuration.
Chapter Outline

9.1 Lewis Electron Dot Structures
9.2 Molecular Geometry
9.3 Polarity and Intermolecular Forces
9.4 Hybridization of Atomic Orbitals
9.5 Covalent Bonding
Lesson Quiz

1. One of the following is not a polyatomic ion
   1. SO₂
   2. SO₃⁻
   3. SO₄²⁻
   4. HSO₃⁻

2. Atoms have ______ potential energy when they are bonded than when they are isolated
   1. higher
   2. lower
   3. equal
   4. variable

3. A bond in which two atoms share one or more pairs of electrons is called a
   1. coordinate bond
   2. coordinate covalent bond
   3. covalent bond
   4. cohesive bond

4. The Lewis structure represents a compound with

   \[
   \begin{array}{c}
   H \quad H \\
   H : \overset{\dddot{}}{\overset{\dddot{}}{\overset{\dddot{}}{\dddot{}}}} : \overset{\dddot{}}{\overset{\dddot{}}{\dddot{}}}: \overset{\dddot{}}{\dddot{}} : H \\
   H \quad H
   \end{array}
   \]

   1. seven single bonds and one double bond
   2. five single bonds and two double bonds
   3. six single bonds and one double bond
   4. eight single bonds

5. A resonance structure
   1. gives a realistic picture of shifting bonds
   2. accurately shows the differing structural isomers
   3. is a hybrid of all the possible structures
   4. shows different atomic charges that exist

6. One of the following does not represent an exception to the octet rule.
9.1. Lewis Electron Dot Structures

1. NO₂
2. PCl₅
3. BH₃
4. BeCl₂

7. The amount of energy needed to break a halogen-halogen bond is
   1. fairly small
   2. high
   3. very high
   4. moderate

8. Which is the proper way to represent the Lewis structure for a polyatomic ion?
   1. [L⁺]
   2. L⁺
   3. [L]⁺
   4. L[⁺]

9. In the CO molecule, the coordinate covalent bond is formed by
   1. carbon contributing an s electron
   2. oxygen contributing a lone pair.
   3. carbon contributing a lone pair
   4. oxygen contributing an s electron

10. A double bond is formed by two atoms
   1. sharing one pair of electrons
   2. sharing two lone pairs of electrons
   3. sharing two pairs of electrons
   4. sharing one lone pair of electrons

True/False:

11. _____ The central atom in a Lewis structure is the least electronegative.
12. _____ A covalent bond forms when two singly occupied orbitals overlap with each other.
13. _____ The structure C≡C is the Lewis structure for a double bond.
14. _____ The electrons that form a covalent bond must have opposite spin.
15. _____ A carbon-carbon double bond has a higher bond energy than a carbon-carbon single bond.
16. _____ The nitrogen-nitrogen triple bond is very reactive.
17. _____ The d-sublevel electrons can be used in bonding in some cases.
18. _____ Many polyatomic ions exhibit resonance forms.
19. _____ In ozone, the two oxygen-oxygen bonds are identical.
20. _____ The F₂ molecule contains two shared pairs of electrons.

Fill in the Blank:

21. Draw the Lewis structures for
   1. OH⁻
   2. NH₃
   3. Cl-Br

22. Draw the structural formulas for

   : Cl : Cl :
23. Which exception to the octet rule is shown by BH3. Explain your answer
24. The two identical bonds in O3 are stronger and shorter than a typical O-O single bond but longer and weaker than an O-O double bond. What can we conclude about the structure of the ozone molecule?
25. How is a resonance structure a type of model for a molecule?
26. Why is the O atom represented by

\[
\cdot\bar{\text{O}}\cdot
\]

in Lewis structures?
27. A structural formula uses ______ to show bonds between atoms.
28. ____________ is the energy required to break a covalent bond between two atoms.

## Answer Key

1. a
2. b
3. c
4. c
5. c
6. d
7. a
8. c
9. b
10. c
11. true
12. true
13. false
14. true
15. true
16. false
17. true
18. true
19. true
20. false

\[
\begin{align*}
\text{H} : \bar{\text{O}} & \quad \left[ \text{H} : \bar{\text{O}} \right]^- \\
\text{H} : \bar{\text{N}} & \quad \text{H} \\
: \bar{\text{Cl}} & \quad \bar{\text{Br}}
\end{align*}
\]
1. Cl-Cl
2. S=C=S
3. O-S-O

21. incomplete octet, B only has three valence electrons. BCl$_3$ gives B a total of six electrons in the outer shell.
22. The ozone molecule has two bonds that are intermediate between a single bond and a double bond.
23. Resonance structures give an idea of what the molecule looks like, but does not give an accurate picture.
24. The two dots (one on either side of the atom) represent the unpaired electrons that can be involved in forming a covalent bond.
25. dashes
26. bond energy
Lesson Quiz

Name ____________________ Class ______________ Date ________

Note: all images in this lesson are in the public domain or taken from CK-12 materials.

Multiple Choice:

1. The VSEPR model explains
   1. reactivity of molecules
   2. molecular composition
   3. spacing of atoms in a molecule
   4. orbital structure of electrons

2. The AB₃ configuration is
   1. trigonal pyramidal
   2. trigonal planar
   3. trigonal parallel
   4. trigonal bipyramidal

3. Sulfur hexafluoride has the following VSEPR configuration
   1. trigonal bipyramidal
   2. tetrahedral
   3. linear
   4. octahedral

4. The AB₂E designation can be seen in
   1. methane
   2. ammonia
   3. ozone
   4. water

5. A distorted tetrahedron molecular geometry is seen in the _______ molecule.
   1. SF₅
   2. ClF₃
   3. SF₄
   4. Cl₂F₅

6. A tetrahedral electron domain geometry can have one of the following molecular geometries
   1. seesaw
   2. T-shaped
   3. square planar
   4. bent

7. Xenon tetrafluoride has a Lewis structure consisting of
   1. four single bonds and two lone pairs
2. four single bonds and one lone pair
3. four single bonds
4. four single bonds and three lone pairs

8. The triiodide ion has a _______ geometry
   1. octahedral
   2. linear
   3. square pyramidal
   4. T-shaped

9. The H-O-H bond angle is
   1. 109.5°
   2. 107°
   3. 104.5°
   4. 103.2°

10. A characteristic of the ClF₃ molecule is
    1. two lone pair sets of electrons
    2. one lone pair set of electrons
    3. trigonal pyramidal domain geometry
    4. tetrahedral domain geometry

True/False:

11. _____ The VSEPR model is based on the fact that electrons repel one another.
12. _____ In the VSEPR model, double and triple bonds are treated differently than single bonds.
13. _____ In the VSEPR model, the central atom is the least electronegative.
14. _____ Only central atom lone pair electrons are considered to affect geometry.
15. _____ Only bonding pair electrons are considered in the electron domain geometry.
16. _____ Sulfur hexafluoride (SF₆) has no lone pair electrons.
17. _____ In a perspective drawing of a molecule, a solid triangle is visualized as receding into the page.
18. _____ All four atoms in BF₃ lie in the same plane.
19. _____ Molecular geometry deals with the three-dimensional arrangement of atoms in a molecule.
20. _____ The valence orbital is the outermost occupied space in an atom.

Fill in the Blank:

21. Name the molecule below. Indicate the VSEPR classification. What does the dotted line represent? What does the solid triangle represent? What does the single solid line represent?

22.

23. Why does the ozone molecule take on a bent configuration?

24.

25. The molecular geometry of ammonia is _______.
26. Why does the sulfur tetrafluoride molecule have a distorted tetrahedron configuration?
27. 

28. Chlorine trifluoride has a _________molecular geometry and _______ lone pairs.
29. The water molecule has ______ lone pairs of electrons on the O atom and a _________ molecular geometry.
30. Name the following types of structures (E = lone pair electrons)

31. Match the structure with the compound

1. PCl$_5$
2. SF$_6$
3. H$_2$O

32. Image 1

33. Image 2

34. Image 3
Answer Key

1. c
2. b
3. d
4. c
5. c
6. d
7. a
8. b
9. c
10. a
11. true
12. false
13. true
14. true
15. false
16. true
17. false
18. true
19. true
20. false
21. Boron trifluoride. Trigonal planar. The dotted line represents a bond receding into the plane of the page, the solid triangle designates a bond coming out of the plane of the page, and the line is for a bond that is in the plane of the page.
22. The central O atom has two lone pairs of electrons that force the outer O atoms into the bent configuration.
23. trigonal pyramidal
24. The central S atom contains a lone pair set of electrons that forces the F atoms out of the plane.
25. T-shaped, two
26. two, bent
   1. trigonal pyramidal
   2. tetrahedral
   3. square pyramidal
Lesson Quiz

Multiple Choice:

1. A bond that shows strong attraction of electrons by one atom is called
   1. polar covalent
   2. non-polar covalent
   3. ionic
   4. covalent

2. A partial charge is represented by
   1. $\Delta$
   2. $\lambda$
   3. $\delta$
   4. $\alpha$

3. The Si-Br bond is
   1. mostly ionic
   2. partially ionic
   3. polar covalent
   4. mostly covalent

4. A dipole is represented by
   1. plus — minus
   2. plus — plus
   3. minus — plus
   4. minus — minus

5. Random motions of electrons are called
   1. London dispersion forces
   2. Lewis dispersion forces
   3. Gibbs dispersion forces
   4. Pauling dispersion forces

6. An atom that could form a H bond with an appropriate H atom is
   1. C
   2. Si
   3. I
   4. O

7. The physical state of an ionic compound at room temperature is
   1. solid
   2. liquid
   3. gas
   4. liquid or gas
8. Hydrogen bonding is a particularly strong form of
   1. ion-dipole interaction
   2. ion-ion interaction
   3. ion-covalent interaction
   4. dipole-dipole interaction

9. The relative ability of an atom to attract electrons is called
   1. electronegativity
   2. electropolarity
   3. electropositivity
   4. electrodispersion

10. One of the following is not a diatomic element
    1. hydrogen
    2. helium
    3. nitrogen
    4. iodine

True/False:
11. _____ A single water molecule can form H-bonds with more than one other water molecule.
12. _____ Hydrogen bonds make it possible for water to be a liquid at room temperature.
13. _____ CsO would be expected to be a mostly covalent compound.
14. _____ Bonds between two non-metal atoms are usually covalent in nature.
15. _____ HI molecules are held together by hydrogen bonds.
16. _____ The Si-I bond would be considered to be polar covalent.
17. _____ Electrons are not shared equally in a non-polar covalent bond.
18. _____ Partial positive and negative charges are used to indicate uneven electron distribution in a covalent bond.
19. _____ A molecule with two poles is called a dipole.
20. _____ When placed between oppositely charged plates, polar molecules orient themselves so that their positive ends are closer to the positive plate.

Fill in the Blank:
21. Classify the following bonds as mostly ionic, polar covalent, or mostly covalent.

   1. Si-P
   2. Ca-O
   3. H-Br

22. Define the following terms:

   1. van der Waals forces
   2. diatomic element
   3. dipole
   4. polar molecule

23. Which of the following pairs of molecular components will form hydrogen bonds? X = remainder of molecule.

   1. X-N-H and H-C-X
   2. X-O-H and H-F
   3. X-O-H and H-N-X
   4. I-H and H-O-X

24. Arrange the following in terms of increasing polarity, beginning with the least polar:
1. Ca-As
2. Pt-I
3. Si-F
4. Al-Cl
5. H-O

25. Can an ether (generic structure C-O-C) form H-bonds with water?

Answer Key

1. a
2. c
3. c
4. a or c
5. a
6. d
7. a
8. d
9. a
10. b
11. true
12. true
13. false
14. true
15. false
16. true
17. false
18. true
19. true
20. false

1. mostly covalent
2. mostly ionic
3. polar covalent

1. The weakest intermolecular force and consist of dipole-dipole forces and dispersion forces.
2. Elements whose natural form is of a diatomic molecule.
3. A molecule with two poles.
4. A molecule in which one end of the molecule is slightly positive, while the other end is slightly negative.

21. b and c
22. least polar Pt-I Ca-As H-O Al-Cl Si-F most polar
23. Ethers have relatively low boiling points due to their inability to form hydrogen bonds with each other.
9.4 Hybridization of Atomic Orbitals

Lesson Quiz

Name___________________ Class________________ Date ________

Multiple Choice:

1. Unpaired electrons are required in order to form
   1. p orbitals
   2. ions
   3. covalent bonds
   4. dipoles

2. In forming a bond, potential energy reaches a minimum
   1. before the electrons interact
   2. at twice the atomic radius
   3. when the nuclei connect
   4. when the bond length is achieved

3. Covalent bonds are formed when
   1. partially filled atomic orbitals overlap
   2. full orbitals interact
   3. partially filled orbitals dissociate
   4. full orbitals hybridize

4. The bond angles in methane are
   1. 107°
   2. 109.5°
   3. 104.5°
   4. 90°

5. The ammonia molecule displays _____ hybridization
   1. sp
   2. sp²
   3. sp³
   4. sp⁴

6. The electron domain geometry for BF₃ leads to _____ hybridization.
   1. sp
   2. sp²
   3. sp³
   4. sp⁴

7. The electron domain geometry of sp² hybrid orbitals is
   1. trigonal planar
   2. linear
3. hexagonal
4. bent

8. One of the following does not have linear electron domain geometry.
   1. BeH₂
   2. NH₃
   3. CO₂
   4. C₂H₂

9. Octahedral electron domain geometry is seen in _____ hybridization.
   1. sp²
   2. sp³d
   3. sp³d²
   4. sp³

10. Ethyne is a linear carbon molecule which contains _____.
    1. two single bonds between the two carbons
    2. a triple bond between the two carbons
    3. a double bond and a single bond between the two carbons
    4. two double bonds between the two carbons

True/False:

11. _____ Overlapping orbitals can be of different types.
12. _____ Hybridization can involve f orbitals.
13. _____ In C=C, one bond is a sigma bond and the other is a pi bond.
14. _____ A sigma bond is formed by overlapping orbitals side to side.
15. _____ In a molecule with linear electron domain geometry, the hybridization of the central atom is sp.
16. _____ If the hybridization of the central atom is sp², the electron domain geometry is trigonal planar.
17. _____ Oxygen in compounds has two lone pair sets of electrons.
18. _____ The pi bond extends above and below the plane of the molecule.
19. _____ Triple bonds are composed of three sigma bonds.
20. _____ In a pi bond, orbitals overlap in a side-to-side fashion.

Fill in the Blank:

21. Define the following terms:
   1. hybridization
   2. valence bond theory
   3. hybrid orbitals
   4. sigma bond
   5. pi bond

22. Describe the bonding that takes place in each of the following diatomic molecules
    1. CS₂
    2. I₂
    3. H₂S

23. Indicate the hybridization of the central carbon atom in each of the following:
    1. CH₄
24. How can SF₆ involve d orbitals when there are no d electrons in either atom?

**Answer Key**

1. c  
2. d  
3. a  
4. b  
5. c  
6. b  
7. a  
8. b  
9. c  
10. b  
11. true  
12. false  
13. true  
14. false  
15. true  
16. true  
17. true  
18. true  
19. false  
20. true

1. The mixing of the atomic orbitals in an atom to produce a set of hybrid orbitals.  
2. Covalent bonds are formed by the overlap of partially filled atomic orbitals.  
3. The atomic orbitals obtained when two or more nonequivalent orbitals from the same atom combine in preparation for bond formation.  
4. A bond formed by the overlap of orbitals in an end-to-end fashion, with the electron density concentrated between the nuclei of the bonding atoms.  
5. A bond formed by the overlap of orbitals in a side-by-side fashion, with the electron density concentrated above and below the plane of the nuclei of the bonding atoms.

1. S=C=S  
2. I-I  
3. H-C≡N

1. sp³, each C forms sigma bonds with the four H atoms
2. \( sp^2 \) C forms sigma bonds with the 2 H atoms. The carbon-oxygen bond is a double bond comprised of an end-to-end sigma bond and a pi bond that overlaps above and below the plane of the C-O connection.

3. \( sp \) The H-C bond is a sigma bond and the C-N bond is a sigma bond plus two pi bonds oriented at right angles to one another.

21. Some p electrons are elevated to d orbital energy levels and occupy those empty spaces.
9.5 Covalent Bonding

Chapter Test

Name ___________________ Class ______________ Date ________

Multiple Choice:

1. In the NH$_4^+$ ion, the coordinate covalent bond is formed by
   1. nitrogen contributing an $s$ electron
   2. nitrogen contributing a lone pair.
   3. hydrogen contributing a lone pair
   4. hydrogen contributing an $s$ electron

2. One of the following represents an exception to the octet rule
   1. F$_2$
   2. C$_2$H$_4$
   3. NO$_2$
   4. O$_3$

3. The AB$_2$ configuration is
   1. linear
   2. trigonal planar
   3. square planar
   4. tetrahedral

4. Methane has the following configuration
   1. trigonal bipyramidal
   2. tetrahedral
   3. linear
   4. octahedral

5. The H-C-H bond angle is
   1. 109.5°
   2. 107°
   3. 104.5°
   4. 103.2°

6. The –N-H —- O-C interaction is an example of
   1. ion-dipole interaction
   2. ion-ion interaction
   3. ion- covalent interaction
   4. dipole-dipole interaction

7. One of the elements below is diatomic
   1. iron
   2. magnesium
3. helium
4. iodine

8. One of the following is not true about van der Waals forces
   1. ion-ion interactions
   2. dipole-dipole forces
   3. dispersion forces
   4. hydrogen bonds

9. CHBr₃ displays ______ hybridization
   1. sp
   2. sp²
   3. sp³
   4. sp⁴

10. A C=C bond consists of
    1. two σ bonds
    2. two π bonds
    3. one π bond
    4. one σ bond and one π bond

True/False:

11. _____ The central atom in a Lewis structure is the most electronegative.
12. _____ The f-sublevel electrons can be used in bonding in some cases.
13. _____ In the VSEPR model, double and triple bonds are treated the same as single bonds.
14. _____ Lone pair electrons on any atom in the structure affect geometry.
15. _____ Carbon dioxide is a linear molecule.
16. _____ In the VSEPR model, valence electron pairs are attracted to one another.
17. _____ Bonds between two non-metal atoms are usually dipole in nature.
18. _____ The symbol δ⁺ is used to indicate a partial positive charge on an atom.
19. _____ Nitrogen in compounds has one lone pair set of electrons.
20. _____ A pi bond is formed by overlapping orbitals side to side.

Fill in the Blank:

21. Which exception to the octet rule is seen in NO₂. Explain your answer.
22. Draw the Lewis structure for
    1. O₂
    2. HBr
    3. F₂

23. Name the molecular geometry structure below

1.

2.

3.
24. Arrange the following in terms of polarity, beginning with the least polar
   1. In-Cl
   2. H-Br
   3. Zn-F
   4. N-O
   5. Ga-Br

25. Define the following terms:
   1. London dispersion forces
   2. diatomic element
   3. coordinate covalent bond
   4. valence shell
   5. valence bond theory

26. Describe the hybridization of the central atom in each of the following:
   1. BF$_3$
   2. CCl$_4$
   3. CO$_2$

---

**Answer Key**

1. b
2. c
3. a
4. b
5. a
6. d
7. d
8. a
9. c
10. d
11. false
12. false
13. true
14. false
15. true
16. false
17. false
18. true
19. true
20. true
21. NO$_2$ The total number of valence electrons is an odd number so all atoms cannot satisfy the octet rule.

1. \[
   \begin{align*}
   & \ddash\ddash O \cdash\cdash O \\
   \end{align*}
   \]
2. \[
   H \cdash\cdash Br
   \]
3. \[
   \ddash\ddash F \cdash\cdash F
   \]

1. trigonal bipyramidal
2. octahedral
3. T-shaped

22. least polar N-O H-Br Ga-Br In-Cl Zn-F most polar

1. Intermolecular forces that occur between all atoms and molecules due to the random motion of electrons
2. Elements whose natural form is of a diatomic molecule
3. A covalent bond in which one of the atoms contributes both of the electrons in the shared pair.
4. The outermost occupied shell of electrons in an atom
5. Covalent bonds are formed by the overlap of partially filled atomic orbitals.

1. sp²
2. sp³
3. sp
CHAPTER 10

The Mole Assessments

Chapter Outline

10.1 THE MOLE CONCEPT
10.2 MASS, VOLUME, AND THE MOLE
10.3 CHEMICAL FORMULAS
10.4 THE MOLE
Lesson Quiz

Name___________________ Class________________ Date________

Multiple Choice:

1. The molecule CH₂CHCH₃CH₂CHF(CH₃)₂CCH₃ contains ___ H atoms
   1. 16
   2. 22
   3. 18
   4. 20

2. An exact quantity
   1. has been measured very carefully
   2. is found by counting
   3. is determined by several labs
   4. has been defined by an international commission

3. A representative particle of silicon dioxide is
   1. a grain of sand
   2. an atom of silicon
   3. a molecule of SiO₃
   4. a molecule of SiO₂

4. One of the following statements about atomic mass is untrue. Atomic mass is
   1. a relative number
   2. based on carbon-12
   3. based on H-1
   4. used to determine molar mass

5. There are _____ chocolate chips in five dozen chocolate chips
   1. 5
   2. 60
   3. 50
   4. 30

6. The _____ is the SI unit for the amount of a substance.
   1. milligram
   2. milliliter
   3. mole
   4. micron

7. The representative particle for O₂ is the
   1. atom
   2. molecule
3. formula unit
4. mole

8. Molar mass is defined as the mass of
   1. one mole of a solution
   2. one mole of representative particles of a substance
   3. $6.02 \times 10^{23}$ particles
   4. one Avogadro’s number of atoms

9. The representative particle for KBr is the
   1. atom
   2. formula unit
   3. molecule
   4. extended lattice

10. The molar mass of an element is
   1. the atomic mass expressed in grams
   2. the molecular mass
   3. the atomic mass expressed in amu
   4. the atomic mass expressed in milligrams

True/False:

11. _____ Dimensional analysis uses conversion factors.
12. _____ The representative particle for H$_2$ is the atom.
13. _____ “Formula mass” can apply to both molecules and ions.
14. _____ Ionic compounds exist as discrete molecules.
15. _____ A mole of CH$_3$CHCH$_3$CH$_2$CHF(CH$_3$)$_2$CCH$_3$ contains nine moles of carbon.
16. _____ Avogadro’s number has been experimentally determined.
17. _____ Atomic masses on the periodic table are usually whole numbers.
18. _____ The molecular mass of a compound is the mass of one mole of that compound.

Short Answers:

19. You are buying chocolate chip cookies for the chess club. There are ten members of the club and they will each eat eight cookies. How many cookies do you need to buy?
20. Convert the given number of particles to moles
   1. $1.7 \times 10^{14}$ atoms of Pt
   2. $3.8 \times 10^{26}$ molecules of ethane
   3. $5.1 \times 10^{12}$ formula units of LiBr.
   4. 3.7 molecules N$_2$
21. Convert the moles to the number of representative particles
   1. 24 mol gold
   2. 0.15 mol trinitrotoluene
   3. 0.57 mol MgBr$_2$
   4. 0.83 mol O$_2$
22. Calculate the formula mass for each of the following compounds
   1. PtCl$_6$
   2. N$_2$
   3. CH$_3$CH$_2$OCH$_3$
   4. FeO$_2$C$_6$H$_4$
1. c
2. b
3. d
4. c
5. b
6. c
7. b
8. b
9. b
10. a
11. true
12. false
13. true
14. false
15. true
16. true
17. false
18. false
19. 10 members \times 8 \text{ cookies/member} = 80 \text{ cookies}

1. \(1.7 \times 10^{14} \text{ atoms} \div 6.02 \times 10^{23} \text{ atoms/mol} = 2.8 \times 10^{-10} \text{ moles}\)
2. \(3.8 \times 10^{26} \div 6.02 \times 10^{23} = 6.31 \times 10^2 \text{ moles}\)
3. \(5.1 \times 10^{12} \div 6.02 \times 10^{23} = 8.47 \times 10^{-12} \text{ moles}\)
4. \(3.7 \div 6.02 \times 10^{23} = 6.15 \times 10^{-24} \text{ moles}\)

1. \((2.4 \text{ mol})(6.02 \times 10^{23} \text{ particles/mol}) = 1.44 \times 10^{25} \text{ particles}\)
2. \((0.15)(6.02 \times 10^{23}) = 9.0 \times 10^{22} \text{ particles}\)
3. \((0.57)(6.02 \times 10^{23}) = 3.43 \times 10^{23} \text{ particles}\)
4. \((0.83)(6.02 \times 10^{23}) = 5.0 \times 10^{23} \text{ particles}\)

1. \(195.08 + 6(35.5) = 408.08 \text{ g/mol}\)
2. \(2 \times 14.01 = 28.02 \text{ g/mol}\)
3. \(3(12) + 8(1.01) + 16.0 = 60.08 \text{ g/mol}\)
4. \(55.85 + 2(16.0) + 6(12.0) + 4(1.01) = 115.9 \text{ g/mol}\)
10.2 Mass, Volume, and the Mole

Lesson Quiz

Name___________________ Class________________ Date________

Multiple Choice:

1. One of the following is not a representative particle
   1. atoms
   2. ions
   3. molecules
   4. formula units

2. In order to convert between mass and number of particles, a conversion to ___ is first required.
   1. grams
   2. volume
   3. moles
   4. atoms

3. The molar mass of CuCl₂ is
   1. 134.6
   2. 98.9
   3. 207.7
   4. 63.55

4. The compound whose molar mass is 212.4 is
   1. CsBr
   2. RbI
   3. KF
   4. CaF₂

5. One of the following is not a property of a gas
   1. compressibility
   2. crowded molecules
   3. negligible particle size
   4. expandability

6. STP stands for
   1. sequentially timed pressure
   2. standard telemetry process
   3. standard temperature and pressure
   4. standard temperature processing

7. Gas density can be calculated using
   1. \( D = \text{molar mass} \times 22.4 \)
   2. \( D = \frac{\text{molecular mass}}{22.4} \)
3. \( D = \text{molecular mass} \times 22.4 \)
4. \( D = \text{molar mass} \div 22.4 \)

8. The mole road map cannot be used in calculations of
   1. volume
   2. concentration
   3. mass
   4. representative particles

9. One mole of gas = 22.4 L only at
   1. 0ºC and 1 atm pressure
   2. 32ºC and one atm pressure
   3. 0ºK and one atm pressure
   4. 0ºF and one atm pressure

10. One of the following does not affect gas volume
    1. high pressure
    2. size of gas molecules
    3. high temperature
    4. low pressure

True/False:

11. _____ The number of moles of a solid can be determined by measuring the volume of the solid.
12. _____ As the temperature increases, the gas volume increases.
13. _____ Units for gas density are kg/L
14. _____ The density for hydrogen is greater than the density of air.
15. _____ A balance can be used to directly determine the number of moles in a material.
16. _____ Equal volumes of H\(_2\) and Xe at STP contain different numbers of particles.
17. _____ Equal volumes of H\(_2\) and Xe at STP have different masses.
18. _____ The density of a gas at a specific temperature and pressure is dependent on its molar mass.

Short Answers:

19. Find the mass of the following materials:
    a. 0.45 moles C
    b. 2.67 moles Hg
    c. 1.9 moles MgCl\(_2\)
    d. 4.3 moles CH\(_2\)=CH-CH\(_2\)F

20. Determine the number of moles in each of the following:
    a. 153 g P\(_2\)O\(_5\)
    b. 75.3 g N\(_2\)O\(_2\)
    c. 17.4 g HCl
    d. 92.6 g CH \equiv C – CHBr\(_2\)

21. How many moles of gas in each of the following?
    a. 839 L of argon
    b. 164 L of hydrogen
    c. 63.7 L of fluorine

22. How much volume is occupied by the following gases?
    a. 26.4 moles sulfur dioxide
b. 14.5 moles helium
c. 32.7 moles chlorine

23. Calculate the density of each of the following gases:
   a. CH₃CH₃
   b. CH ≡ CH
   c. H₂S

24. Calculate the molar mass of the gas when given the density at STP
   a. density = 2.86 g/L
   b. density = 1.96 g/L
   c. density = 0.09 g/L

---

**Answer Key**

1. b
2. c
3. a
4. a
5. b
6. c
7. d
8. b
9. a
10. b
11. false
12. true
13. false
14. false
15. false
16. false
17. true
18. true

**Calculations**

1. \((12.0 \text{ g/mole})(0.45 \text{ moles}) = 5.4 \text{ g}\)
2. \((200.6 \text{ g/mole})(2.67 \text{ moles}) = 535.6 \text{ g}\)
3. \(Mg 1 \times 24.3 = 24.3\)
4. \(Cl 2 \times 35.5 = 71.0\)
5. \(= 95.3 \text{ g/mol}\)
6. \((95.3 \text{ g/mol})(1.9 \text{ moles}) = 181.1 \text{ g}\)
7. \(C 3 \times 12.0 = 36.0\)
8. \(H 5 \times 1.01 = 5.05\)
9. \(f 1 \times 19.0 = 19.0\)
10. \(= 60.1 \text{ g/mole}\)
11. \((60.1 \text{ g/mole})(4.3 \text{ moles}) = 258.2 \text{ g}\)
a. 
\[ P \times 31.0 = 62.0 \]  
\[ O \times 16.0 = 80.0 \]  
\[ = 142.0 \text{ g/mole} \]

b. \( 153 \text{ g} \div 142.0 \text{ g/mole} = 1.08 \text{ moles} \)

c. 
\[ N \times 14.01 = 28.02 \]  
\[ O \times 16.0 = 32.0 \]  
\[ = 60.02 \text{ g/mole} \]

d. \( 75.3 \text{ g} \div 60.02 \text{ g/mole} = 1.25 \text{ moles} \)

e. 
\[ H \times 1.01 = 1.01 \]  
\[ Cl \times 35.5 = 35.5 \]  
\[ = 36.51 \text{ g/mole} \]

f. \( 17.4 \text{ g} \div 36.51 \text{ g/mole} = 0.49 \text{ moles} \)

g. 
\[ C \times 12.0 = 36.0 \]  
\[ H \times 1.01 = 2.02 \]  
\[ Br \times 79.9 = 159.8 \]  
\[ = 197.8 \text{ g/mole} \]

h. \( 92.6 \text{ g} \div 197.8 \text{ g/mole} = 0.46 \text{ moles} \)

1. \( 839 \text{ L} \div 22.4 \text{ L/mole} = 37.5 \text{ moles} \)
2. \( 164 \text{ L} \div 22.4 \text{ L/mole} = 7.32 \text{ moles} \)
3. \( 63.7 \text{ L} \div 22.4 \text{ L/mole} = 2.84 \text{ moles} \)

1. \( (26.4 \text{ moles})(22.4 \text{ L/mole}) = 591.4 \text{ L} \)
2. \( (14.5 \text{ moles})(22.4 \text{ L/mole}) = 324.8 \text{ L} \)
3. \( (32.7 \text{ mole})(22.4 \text{ L/mole}) = 732.5 \text{ L} \)

a. 
\[ C \times 12.0 = 24.0 \]  
\[ H \times 1.01 = 6.06 \]  
\[ = 30.06 \text{ g/mole} \]

b. \( 30.06 \text{ g/mole} \div 22.4 \text{ L/mole} = 1.34 \text{ g/L} \)

c. 
\[ C \times 12.0 = 24.0 \]  
\[ H \times 1.01 = 2.02 \]  
\[ = 26.02 \text{ g/mole} \]

d. \( 26.02 \text{ g/mole} \div 22.4 \text{ L/mole} = 1.16 \text{ g/L} \)
e. 

\[
\begin{align*}
H \times 2 \times 1.01 &= 2.02 \\
S \times 32.1 &= 32.1 \\
&= 34.12 \text{ g/mole}
\end{align*}
\]

f. 34.12 g/mole ÷ 22.4 L/mole = 1.52 g/L

1. \((2.86 \text{ g/L})(22.4 \text{ L/mole}) = 64.1 \text{ g/mole}\)
2. \((1.96 \text{ g/L})(22.4 \text{ L/mole}) = 43.9 \text{ g/mole}\)
3. \((0.09 \text{ g/L})(22.4 \text{ L/mole}) = 2.02 \text{ g/mole}\)
Lesson Quiz

Name___________________ Class______________ Date________

Multiple Choice:

1. A hydrate can be converted to its anhydrous form by
   1. heating
   2. vacuum distillation
   3. chemical treatment
   4. lowering the temperature

2. The percent composition tells us
   1. the purity of the compound
   2. the amount of each element present
   3. the yield of the reaction
   4. the amount needed for a reaction

3. An empirical formula will not provide the following information
   1. lowest whole-number ratio of elements
   2. elemental composition of the compound
   3. the molecular mass of the compound
   4. the relative amounts of elements in the compound

4. Mass spectrometers can determine
   1. molecular mass
   2. molecular formula
   3. molar mass
   4. molar formula

5. Percent composition is defined as
   1. mass of element \times mass of compound \times 100
   2. mass of compound \div mass of element \times 100
   3. (mass of element \div mass of compound) \times 100
   4. (mass of element \div mass of compound) \div 100

6. One of the following is not needed to determine percent of a hydrate’s mass
   1. number of water molecules
   2. molecular mass of compound
   3. reactivity of compound
   4. molecular mass of water

7. The molecular formula for cyclohexane is \text{C}_6\text{H}_{12}. The correct empirical formula is
   1. \text{C}_2\text{H}_4
   2. \text{CH}_3
3. CH
4. CH₂

8. Cobalt chloride hydrate has the following color
   1. magenta
   2. blue
   3. orange
   4. yellow

True/False:

9. _____ Purity of the compound affects determination of percent composition.
10. _____ Hydrate water molecules are bound to the crystal lattice of a formula unit.
11. _____ Ionic compounds exist as discrete molecules.
12. _____ Elemental analysis gives information for determination of empirical formulas.
13. _____ In many situations, the molecular formula and the empirical formula are the same.
14. _____ Acetic acid and glucose have different empirical formulas.
15. _____ The empirical formula mass is the molar mass represented by the empirical formula.
16. _____ Hydrates may contain fractional molecules of water.

Short Answers:

17. Determine the percent composition of the following compounds.
   a. RbI
   b. Na₂O
   c. CH₃OH

18. Determine the percent composition from the following analytical data:
   a. Analysis of a 30.9 g sample of a compound yielded 9.63 g S and 21.3 g Cl.
   b. A sample with a mass of 95.2 g contains 40.8 g C and 54.4 g O.
   c. A sample has a mass of 0.432 g and contains only O and F. The oxygen content is 0.128 g.

19. Calculate the percent by mass of water in the following compounds:
   a. LiNO₃·₃H₂O
   b. CoF₂·₄H₂O
   c. ZnSO₄·₇H₂O

20. Calculate the empirical formula from the data below:
   a. 47.9% Zn, 52.1% Cl
   b. 20.23% Al, 79.77% Cl
   c. 24.74% K, 34.7% Mn, 40.50% O

21. Determine the molecular formula:
   a. empirical formula ClCH₂ molar mass 98.96 g/mol
   b. empirical formula C₁₀H₁₂O₂ molar mass 318.31 g/mol
   c. empirical formula C₄H₅N₂ molar mass 194.16 g/mol

Answer Key

1. a
2. b 
3. c 
4. c 
5. c 
6. c 
7. d 
8. a 
9. true 
10. true 
11. false 
12. true 
13. true 
14. false 
15. true 
16. false

a. \( Rb = 85.5 \text{ g/mol} \quad I = 126.9 \text{ g/mol} \quad RbI = 212.4 \text{ g/mol} \)
b. 
\[
\%Rb = \frac{85.5 \text{ g/mol}}{212.4 \text{ g/mol}} \times 100 = 40.3\%
\]
\[
\%I = \frac{126.9 \text{ g/mol}}{224 \text{ g/mol}} \times 100 = 59.7\%
\]
c. \( 2Na = 46.0 \text{ g/mol} \quad O = 16 \text{ g/mol} \quad Na_2O = 62.0 \text{ g/mol} \)
d. 
\[
\%Na = \frac{46.0 \text{ g/mol}}{62.0 \text{ g/mol}} \times 100 = 74.2\%
\]
\[
\%O = \frac{16 \text{ g/mol}}{62.0 \text{ g/mol}} \times 100 = 25.8\%
\]
e. \( C = 12.0 \text{ g/mol} \quad 4H = 4.04 \text{ g/mol} \quad O = 16.0 \text{ g/mol} \quad CH_3OH = 32.04 \text{ g/mol} \)
f. 
\[
\%C = \frac{12 \text{ g/mol}}{32.04 \text{ g/mol}} \times 100 = 37.4\%
\]
\[
\%H = \frac{4.04 \text{ g/mol}}{32.04 \text{ g/mol}} \times 100 = 12.6\%
\]
\[
\%O = \frac{16.0 \text{ g/mol}}{32.04 \text{ g/mol}} \times 100 = 50.0\%
\]
1. \( %S = (9.3 \text{ g } ÷ 30.9 \text{ g}) \times 100 = 31.1\% \quad %Cl = (21.3 \text{ g } ÷ 30.9 \text{ g}) \times 100 = 68.9\% \)
2. \( %C = (40.8 \text{ g } ÷ 95.2 \text{ g}) \times 100 = 42.9\% \quad %O = (54.4 \text{ g } ÷ 95.2 \text{ g}) \times 100 = 57.1\% \)
3. \( 0.432 \text{ g } ÷ 0.128 \text{ g} = 0.304 \text{ g} F \)
\[ %O = (0.128 \text{ g } ÷ 0.432 \text{ g}) \times 100 = 29.6\% \quad %F = (0.304 \text{ g } ÷ 0.432 \text{ g}) \times 10 = 70.4\% \]
17. mass of one molecule of water: 18.02 g/mole
a. mass of \( LiNO_3 = 6.94 + 14.01 + 3(16.0) = 68.95 \text{ g/mol} \)
b. mass of hydrate = 68.95 + 3(18.02) = 123.01 g/mol
c. \( %\text{water} = (54.06 ÷ 123.01) \times 100 = 43.9\% \)
d. mass of \( CoF_2 = 58.93 + 2(19.00) = 96.93 \text{ g/mol} \)
e. mass of hydrate = 96.93 + 4(18.02) = 169.01 g/mol
f. \( %\text{water} = (72.08 ÷ 169.01) \times 100 = 45.6\% \)
10.3. Chemical Formulas

- **g.** mass of $\text{ZnSO}_4 = 65.39 + 32.07 + 4(16.0) = 161.46 \text{ g/mol}$
- **h.** mass of hydrate $= 161.5 + 7(18.02) = 287.6 \text{ g/mol}$
- **i.** % water $= (126.14 \div 287.6) \times 100 = 43.9\%$

- **a.** $\frac{47.9 \text{ g}}{65.39 \text{ g/mol}} = 0.73 \text{ mol}$  $\frac{52.1 \text{ g}}{35.5 \text{ g/mol}} = 1.47 \text{ mol}$
- **b.** $\text{Zn}^{0.73} \times 1 \text{ Cl}^{1.47} = \text{ZnCl}_2$
- **c.** $\frac{20.23 \text{ g/mol}}{26.98 \text{ g/mol}} = 0.75 \text{ mol}$  $\frac{79.77 \text{ g/mol}}{35.5 \text{ g/mol}} = 2.24 \text{ mol}$
- **d.** $\text{Al}^{0.75} \times 1 \text{ Cl}^{2.24} = \text{AlCl}_3$
- **e.** $\frac{24.74 \text{ g/mol}}{39.1 \text{ g/mol}} = 0.63 \text{ mol}$  $\frac{34.76 \text{ g/mol}}{54.94 \text{ g/mol}} = 0.63$  $\frac{40.5 \text{ g/mol}}{16.0 \text{ g/mol}} = 2.53 \text{ mol}$
- **f.** $\text{K}^{0.63} \times 1 \text{ MN}^{0.63} \times 1 \text{ O}^{2.53} = \text{KMnO}_4$

- **a.** empirical mass $12.0 + 1.01 + 35.5 = 48.51$
- **b.** $\frac{98.96}{48.51} = 2.04$  molecular formula $= \text{ClCH}_2 \times 2 = \text{C}_2\text{H}_4\text{Cl}_2$
- **c.** empirical mass $10(12) + 7(1.01) + 2(16) = 159.07$
- **d.** $\frac{318.31}{159.07} = 2.00$  molecular formula $= \text{C}_{10}\text{H}_7\text{O}_2 \times 2 = \text{C}_{20}\text{H}_{14}\text{O}_4$
- **e.** empirical mass $4(12) + 5(1.01) + 16.0 + 2(14.0) = 97.05$
- **f.** $\frac{194.16}{97.05} = 2.0$  molecular formula $= \text{C}_8\text{H}_{10}\text{O}_2\text{N}_4$
Multiple Choice:

1. A representative particle of carbon dioxide is
   1. an atom
   2. a molecule
   3. 22.4L
   4. an ion

2. Atomic mass is based on
   1. C-12
   2. H-1
   3. O-16
   4. C-14

3. The mole is the SI unit for
   1. volume of material
   2. mass of reactant
   3. amount of a substance
   4. mass of a substance

4. The term amu stands for
   1. atomic molecular unit
   2. atomic mass unit
   3. atomic molar unit
   4. anion mass unit

5. Gas volume is not affected by
   1. high pressure
   2. temperature
   3. size of gas molecules
   4. low pressure

6. One mole of gas equals _____ at STP
   1. 22.4 L
   2. 22.8 L
   3. 23.9 L
   4. 22.4 mL

7. The representative particle for HCl is the
   1. atom
   2. molecule
3. ion
4. formula unit

8. Water molecules in a hydrate are
   1. ionically attached to the molecule
   2. covalently bound to the molecule
   3. covalently bound to the lattice
   4. incorporated into the crystal lattice

9. The amount of each substance present in a compound can best be learned from the
   1. molar mass
   2. empirical formula
   3. percent composition
   4. empirical mass

10. The molecular formula for ethylbenzene is C₈H₁₀. The empirical formula would be
    1. C₃H₅
    2. C₄H₆
    3. C₄H₅
    4. C₃H₆

True/False:

11. _____ Units for gas density are g/mL
12. _____ The representative particle for H₂ is the molecule.
13. _____ Equal volumes of N₂ and Rn gases have the same number of particles.
14. _____ The molar mass of a compound is the mass of one mole of representative particles of that compound.
15. _____ The amount of matter in a gas can be measured by its volume.
16. _____ The term formula mass can only be applied to molecules.
17. _____ The units for molar mass are g/mol.
18. _____ Gas volume increases when the temperature decreases.
19. _____ One atom of the most abundant isotope of hydrogen has a mass of approximately 1 amu.
20. _____ Empirical formulas are determined using elemental analysis.

Short Answers:

21. Make the following calculations of properties of gases
    a. How many moles of SO₂ are in 745 L at STP?
    b. What is the volume of 325 moles of N₂?
    c. What is the density of a gas with a molar mass of 78 g/mole?

22. Determine the following formula masses:
    a. NH₄Cl
    b. NaHCO₃
    c. C₁₇H₁₈F₂NO

23. How many moles are each of the following?
    a. 270 g KBr
    b. 397 g Li₂SO₄

24. Calculate the percent by mass of water in the following compounds
    a. CoCl₂ · 6H₂O
    b. PbCl₂ · 3H₂O
25. How many particles are there in
   a. 0.2 moles FeSO₄
   b. 1.7 moles CH₄

26. What is the mass of
   a. 1.7 moles CaBr₂
   b. 0.48 moles N₂O

27. Determine the percent composition for the following
   a. A compound contains 6.2 g C, 4.1 g H, and 15.9 g O
   b. A compound contains 1.2 moles C and 3.2 moles H

**Answer Key**

1. b
2. a
3. c
4. b
5. c
6. a
7. b
8. d
9. c
10. c
11. false
12. true
13. true
14. true
15. true
16. false
17. true
18. false
19. true
20. true

1. \(745 \text{ L} \div 22.4 \text{ L/mol} = 33.3 \text{ moles}\)
2. \((325 \text{ moles})(22.4 \text{ L/mol}) = 7280 \text{ L}\)
3. \(78 \text{ g/mole} \div 22.4 \text{ L/mole} = 3.48 \text{ g/L}\)

1. \(14.0 + 4(1.01) + 35.5 = 53.5 \text{ g/mole}\)
2. \(23 + 1.01 + 12.0 + 3(16.0) = 84.01 \text{ g/mol}\)

a. \(39.1 + 79.9 = 119.0 \text{ g/mole}\)

b. \(270 \div 119.0 \text{ g/mol} = 2.27 \text{ mole}\)

c. \(2(6.94) + 32.07 + 4(16.0) = 109.95 \text{ g/mole}\)

d. \(397 \div 109.95 \text{ g/mole} = 3.61 \text{ moles}\)

1. \((0.2 \text{ moles})(6.02 \times 10^{23} \text{ particles/mole}) = 1.20 \times 10^{23} \text{ particles}\)
2. \((1.7 \text{ moles})(6.02 \times 10^{23} \text{ particles/mole}) = 1.02 \times 10^{24} \text{ particles}\)

a. \(40.1 + 2(79.9) = 199.9 \text{ g/mole}\)

b. \((1.7 \text{ moles})(199.9 \text{ g/mole}) = 339.8 \text{ g}\)
c. \(2(14.0) + 16 = 44.0 \text{ g/mole}\)
d. \((0.48 \text{ moles})(44.0 \text{ g/moles}) = 21.12 \text{ g}\)

a. \(6.2 + 4.1 + 15.9 = 26.2 \text{ g}\)
b.
\[
\%C = \left(\frac{6.2}{26.2}\right) \times 100 = 23.7\% \\
\%H = \left(\frac{4.1}{26.2}\right) \times 100 = 15.6\%
\]
c. \(1.2 + 3.2 = 4.4 \text{ moles}\)
d.
\[
\%C = \left(\frac{1.2}{4.4}\right) \times 100 = 27.3\% \\
\%H = \left(\frac{3.2}{4.4}\right) \times 100 = 72.7\% 
\]
Chapter 11

Chemical Reactions Assessments

Chapter Outline

11.1 Chemical Equations
11.2 Types of Chemical Reactions
11.3 Chemical Reactions
11.1 Chemical Equations

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. A product is
   1. material acted upon
   2. material made in the reaction
   3. starting material
   4. process of reacting

2. Which of the following is not a word equation?
   1. flour + water + yeast → bread
   2. sodium + water → sodium hydroxide + hydrogen
   3. drive down highway → don’t get lost
   4. silver nitrate + sodium chloride → silver chloride + sodium nitrate

3. The proper form for a chemical equation is
   1. reactants → products
   2. reactants produce products
   3. products → reactants
   4. products are formed from reactants

4. The symbol $\Delta$ in a chemical equation means
   1. change in volume
   2. change in rate
   3. material is heated
   4. material is cooled

5. The symbol $\text{Pt}^\rightarrow$ indicates that
   1. Pt is one of the reagents
   2. Pt is a catalyst in the reaction
   3. Pt is a product
   4. Pt is added after the reaction starts

6. Only one of the following is used to describe the physical state of a reactant or product
   1. w
   2. p
   3. l
   4. r

7. In a chemical formula, the subscript stands for the number of
   1. atoms of an element in the compound
   2. molecules in the reaction
3. products that will be formed
4. reactants needed for the reaction to proceed

8. Balancing equations relies on the law of
   1. conservation of energy
   2. conservation of momentum
   3. conservation of mass
   4. conservation of heat

9. Only one of the following techniques is used in balancing chemical equations
   1. change subscripts
   2. change coefficients
   3. add extra materials
   4. only balance key reagents

10. Which of the following shows the balanced equation for the combustion of ethane(C\textsubscript{2}H\textsubscript{6})?
    1. \( \text{C}_2\text{H}_6 + \text{O}_2 \rightarrow 2\text{CO}_2 + \text{H}_2\text{O} \)
    2. \( \text{C}_2\text{H}_6 + \text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O} \)
    3. \( 2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O} \)
    4. \( 2\text{C}_2\text{H}_6 + 6\text{O}_2 \rightarrow 2\text{CO}_2 + 6\text{H}_2\text{O} \)

True/False:

11. _____ Atoms cannot be created or destroyed.
12. _____ The symbol \((s)\) means the material is in solution.
13. _____ In burning, a material reacts with hydrogen gas.
14. _____ Tarnish is silver sulfide
15. _____ A skeleton equation shows only the important reactants and products.
16. _____ The symbol \(\rightleftharpoons\) represents a reversible reaction.
17. _____ Equations should be balanced even if the formulas are incorrect.
18. _____ In balancing equations, a coefficient of 1 is never shown.
19. _____ All coefficients must be at the lowest possible ratio in a properly balanced equation.
20. _____ A skeleton equation shows the relative amounts of all materials in the reaction.

Fill in the Blank:

21. Write skeleton equations for the following reactions:
    1. barium chloride + potassium sulfate \(\rightarrow\) barium sulfate + potassium chloride.
    2. calcium chloride + potassium phosphate \(\rightarrow\) potassium chloride + calcium phosphate
    3. calcium hydroxide + sulfuric acid \(\rightarrow\) calcium sulfate + water
    4. sodium + iron(III) chloride \(\rightarrow\) iron + sodium chloride

22. Balance the following equations:
    1. \( \text{AgI} + \text{Na}_2\text{S} \rightarrow \text{Ag}_2\text{S} + \text{NaI} \)
    2. \( \text{Ba}_3\text{N}_2 + \text{H}_2\text{O} \rightarrow \text{Ba(OH)}_2 + \text{NH}_3 \)
    3. \( \text{CaCl}_2 + \text{Na}_3\text{PO}_4 \rightarrow \text{Ca}_3(\text{PO}_4)_2 + \text{NaCl} \)
    4. \( \text{FeS} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3 + \text{SO}_2 \)
    5. \( \text{PCl}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{PO}_4 + \text{HCl} \)
    6. \( \text{KClO}_3 + \text{P}_4 \rightarrow \text{P}_4\text{O}_{10} + \text{KCl} \)
    7. \( \text{Fe} + \text{HC}_2\text{H}_3\text{O}_2 \rightarrow \text{Fe(C}_2\text{H}_3\text{O}_2)_3 + \text{H}_2 \)
    8. \( \text{NH}_4\text{OH} + \text{KAl(SO}_4)_2 \cdot 12\text{H}_2\text{O} \rightarrow \text{Al(OH)}_3 + (\text{NH}_4)_2\text{SO}_4 + \text{KOH} + \text{H}_2\text{O} \)
11.1. Chemical Equations

Answer Key

1. b
2. c
3. a
4. c
5. b
6. c
7. a
8. c
9. b
10. c
11. true
12. false
13. false
14. true
15. false
16. true
17. false
18. true
19. true
20. false

1. \( \text{BaCl}_2 + \text{K}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + \text{KCl} \)
2. \( \text{CaCl}_2 + \text{K}_3\text{PO}_4 \rightarrow \text{KCl} + \text{Ca}_3(\text{PO}_4)_2 \)
3. \( \text{Ca}(\text{OH})_2 + \text{H}_2\text{SO}_4 \rightarrow \text{CaSO}_4 + \text{H}_2\text{O} \)
4. \( \text{Na} + \text{FeCl}_3 \rightarrow \text{Fe} + \text{NaCl} \)

1. \( 2\text{AgI} + \text{Na}_2\text{S} \rightarrow \text{Ag}_2\text{S} + 2\text{NaI} \)
2. \( \text{Ba}_3\text{N}_2 + 6\text{H}_2\text{O} \rightarrow 3\text{Ba}(\text{OH})_2 + 2\text{NH}_3 \)
3. \( 3\text{CaCl}_2 + 2\text{Na}_3\text{PO}_4 \rightarrow \text{Ca}_3(\text{PO}_4)_2 + 6\text{NaCl} \)
4. \( 4\text{FeS} + 7\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3 + 4\text{SO}_2 \)
5. \( \text{PCl}_5 + 4\text{H}_2\text{O} \rightarrow 5\text{HCl} + \text{H}_3\text{PO}_4 \)
6. \( 10\text{KClO}_3 + 3\text{P}_4 \rightarrow 3\text{P}_4\text{O}_{10} + 10\text{KCl} \)
7. \( 2\text{Fe} + 6\text{HC}_2\text{H}_3\text{O}_2 \rightarrow 2\text{Fe}(\text{C}_2\text{H}_3\text{O}_2)_3 + 3\text{H}_2 \)
8. \( 4\text{NH}_4\text{OH} + \text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O} \rightarrow \text{Al}(\text{OH})_3 + 2(\text{NH}_4)_2\text{SO}_4 + \text{KOH} + 12\text{H}_2\text{O} \)
Lesson Quiz

1. One of the following is not a type of chemical reaction
   1. combination
   2. decomposition
   3. subtraction
   4. hydrogen replacement

2. The general equation for a single-replacement reaction is
   1. \[ A + B \rightarrow AB \]
   2. \[ A + BC \rightarrow AC + B \]
   3. \[ AC + BD \rightarrow AD + BC \]
   4. \[ AB \rightarrow A + B \]

3. Sr reacts with HCl in a ______ reaction
   1. decomposition
   2. halogen replacement
   3. hydrogen replacement
   4. combination

4. Metal hydroxides form _____ when heated
   1. metal oxide + water
   2. metal + water
   3. metal + oxygen
   4. metal oxide + hydrogen

5. One of the following statements about combination reactions is not true.
   1. magnesium reacts with oxygen to form magnesium oxide
   2. the reaction of nonmetals with each other forms a molecular compound
   3. a combination reaction is also called a retrosynthesis
   4. sodium and chlorine react to form a compound

6. \[ Al + Ni(NO_{3})_{2} \rightarrow Al(NO_{3})_{3} + Ni \] is an example of a ______ reaction
   1. combination
   2. decomposition
   3. double-replacement
   4. single-replacement

7. One of the following is not a possible outcome of a double-replacement reaction
   1. formation of a gas
   2. formation of a macromolecular compound
3. formation of a precipitate
4. formation of a molecular compound

8. A combustion reaction involves a reaction with ____ gas.
   1. N₂
   2. H₂
   3. Cl₂
   4. O₂

9. One of the following does not react with water
   1. Co
   2. Li
   3. Ca
   4. Ba

10. In the reaction between silver nitrate and sodium chloride, _____ is formed.
    1. hydrogen gas
    2. a precipitate
    3. chlorine gas
    4. water

True/False:

11. _____ There are four basic types of reactions.
12. _____ Magnesium is a more reactive metal than copper.
13. _____ Potassium can be stored in open air.
14. _____ Some combination reactions are also combustion reactions.
15. _____ Fluorine exists in nature as a monatomic element.
16. _____ Sulfur can form more than one compound when reacting with oxygen.
17. _____ In a decomposition reaction, a binary compound forms its elements.
18. _____ A metal carbonate decomposes to form a metal oxide and water.
19. _____ Transition metals can often form more than one compound when combined with oxygen.
20. _____ Ca + Br₂ → CaBr₂.

Fill in the Blank:

21. Write the formulas for the product(s) of the following reactions and balance the resulting equations. Write NR if no reaction takes place.
   1. Zn + HCl →
   2. Pb + ZnSO₄ →
   3. CaH₂O₆ + O₂ →
   4. AlCl₃ + H₂Cl₂ →
   5. Ag + S →
   6. Na + H₂O →
   7. AgNO₃ + K₃PO₄ →
   8. Zn + NaF →
   9. F₂ + MgCl₂ →
   10. Mg + O₂ →
   11. Cu(OH)₂ + HBr →
   12. CH₃CH₂OH + O₂ →
Answer Key

1. c
2. b
3. c
4. a
5. c
6. d
7. b
8. d
9. a
10. b
11. false
12. true
13. false
14. true
15. false
16. true
17. true
18. false
19. true
20. false

1. \( \text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + 2\text{H}_2 \)
2. \( \text{Pb} + \text{ZnSO}_4 \rightarrow \text{NR} \)
3. \( \text{C}_6\text{H}_12\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O} \)
4. \( 2\text{AlI}_3 + 3\text{HgCl}_2 \rightarrow 2\text{AlCl}_3 + 3\text{HgI}_2 \)
5. \( \text{sAg} + \text{S} \rightarrow \text{Ag}_2\text{S} \)
6. \( 2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2 \)
7. \( 3\text{AgNO}_3 + \text{K}_3\text{PO}_4 \rightarrow \text{Ag}_3\text{PO}_4 + 3\text{KNO}_3 \)
8. \( \text{Zn} + \text{NaF} \rightarrow \text{NR} \)
9. \( \text{F}_2 + \text{MgCl}_2 \rightarrow \text{MgF}_2 + \text{Cl}_2 \)
10. \( 2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO} \)
11. \( \text{Cu(OH)}_2 + 2\text{HBr} \rightarrow \text{CuBr}_2 + 2\text{H}_2\text{O} \)
12. \( \text{CH}_3\text{CH}_2\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O} \)
Chapter Test

1. The reaction of calcium and oxygen to form calcium oxide is written
   1. \( Ca + O \rightarrow CaO \)
   2. \( Ca + O_2 \) yields \( CaO \)
   3. \( Ca + O_2 \rightarrow CaO_2 \)
   4. \( 2Ca + O_2 \rightarrow 2CaO \)

2. One of the following is not a type of chemical reaction
   1. double-replacement
   2. single-replacement
   3. oxygen replacement
   4. halogen replacement

3. The general equation for a double-replacement reaction is
   1. \( A + B \rightarrow AB \)
   2. \( A + BC \rightarrow AC + B \)
   3. \( AC + BD \rightarrow AD + BC \)
   4. \( AB \rightarrow A + B \)

4. The reaction \( CH_4 + O_2 \rightarrow CO_2 + H_2O \) is both a combination reaction and a ______ reaction.
   1. combustion
   2. double-replacement
   3. single-replacement
   4. decomposition

5. A skeleton equation
   1. shows all the needed materials and their amounts
   2. shows only the needed materials
   3. shows only the important materials
   4. shows only the amounts of needed materials

6. Hydrogen at room temperature would have the following subscript
   1. \( (s) \)
   2. \( (l) \)
   3. \( (g) \)
   4. \( (w) \)

7. \( Ca + O_2 \rightarrow CaO \) violates the law of
   1. conservation of energy
   2. conservation of heat
3. conservation of mass
4. conservation of elements

8. The equation \(2C_8H_{18} + 24O_2 \rightarrow 16CO_2 + 18H_2O\) has an imbalance of ____ atoms
   1. C
   2. O
   3. H
   4. H + O

9. One of the following metals reacts with cold water
   1. Cr
   2. Ba
   3. Ni
   4. Mg

10. The most reactive halogen is
    1. iodine
    2. bromine
    3. fluorine
    4. chlorine

**True/False:**

11. _____ Magnesium reacts dramatically when ignited.
12. _____ Transition metals only form one product in a reaction with oxygen.
13. _____ \(H_2 + O_2 \rightarrow H_2O\) is a skeleton equation.
14. _____ The symbol \((aq)\) represents a material in water-based solution.
15. _____ The symbol \(\Delta\) indicates a catalyst is used in the reaction.
16. _____ A combination reaction can result in the formation of a more complex product.
17. _____ Oxides of sulfur help produce acid rain.
18. _____ The decomposition of HgO produces a change in color.
19. _____ Sodium hydroxide decomposes to form sodium oxide and water.
20. _____ Ni will react with steam, but not cold water.

**Fill In the Blank:**

21. Indicate the type of reaction represented by each equation:
   1. \(Na_3PO_4 + 3KOH \rightarrow 3NaOH + K_3PO_4\)
   2. \(Pb + FeSO_4 \rightarrow PbSO_4 + Fe\)
   3. \(C_6H_{12} + 9O_2 \rightarrow 6CO_2 + 6H_2O\)
   4. \(CaCO_3 \rightarrow CaO + CO_2\)
   5. \(P_4 + 3O_2 \rightarrow 2P_2O_3\)

22. Predict the product(s) of the following reactions.
   1. \(AlBr_3 + Cl_2 \rightarrow\)
   2. \(Mgl_2 + Mn(SO_3)_2 \rightarrow\)
   3. \(RbNO_3 + BeF_2 \rightarrow\)
   4. \(Al + Ag_2CrO_4 \rightarrow\)

23. Balance the following equations:
   1. \(B_2Br_6 + HNO_3 \rightarrow B(NO_3)_3 + HBr\)
   2. \(NH_3 + O_2 \rightarrow NO + H_2O\)
3. \( CO + H_2 \rightarrow C_8H_{18} + H_2O \)
4. \( Na_3PO_4 + HCl \rightarrow NaCl + H_3PO_4 \)
5. \( HClO_4 + P_4O_{10} \rightarrow H_3PO_4 + TCl_2O_7 \)

**Answer Key**

1. d
2. c
3. c
4. a
5. b
6. c
7. c
8. b
9. b
10. c
11. true
12. false
13. true
14. true
15. false
16. true
17. true
18. true
19. true
20. false

1. double-replacement
2. single-replacement
3. combustion or combination
4. decomposition
5. combination

1. \( AlCl_3 + Br_2 \)
2. \( Mg(SO_3)_2 + MnI_2 \)
3. \( RbF + Be(NO_3)_2 \)
4. \( Al_2(CrO_4)_3 + Ag \)

1. \( B_2Br_6 + 6HNO_3 \rightarrow 2B(NO_3)_3 + 6HBr \)
2. \( 4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O \)
3. \( 8CO + 17H_2 \rightarrow C_8H_{18} + 8H_2O \)
4. \( Na_3PO_4 + 3HCl \rightarrow 3NaCl + H_3PO_4 \)
5. \( 12HClO_4 + P_4O_{10} \rightarrow 4H_3PO_4 + 6Cl_2O_7 \)
Chapter 12. Stoichiometry Assessments

Chapter Outline

12.1 Mole Ratios
12.2 Stoichiometric Calculations
12.3 Limiting Reactant and Percent Yield
12.4 Chapter Twelve Exam
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. You invite three other people over to watch the Super Bowl. In preparing for the day, you assume each person will consume three soft drinks and \( \frac{1}{2} \) bag of chips. At the grocery store, you purchase one 12-pack of soft drinks and two bags of chips. However, each of the three brings a friend, requiring an emergency run back to the grocery store. What do you buy to deal with the increase in friends?
   1. one bag of chips
   2. \( \frac{3}{4} \) of a 12-pack of soft drinks
   3. 1.5 bags of chips and \( \frac{3}{4} \) of a 12-pack of soft drinks
   4. two bags of chips and a 12-pack of soft drinks

2. Stoichiometric calculations require
   1. the number of grams of reactant
   2. the number of moles of product
   3. a balanced equation
   4. a word equation

3. Amounts of materials for chemical reactions are generally expressed in all the units below except
   1. moles
   2. \( \mu \)mL
   3. grams
   4. \( \mu \)moles

4. Dinner for six is being prepared. The menu includes broccoli, which one person refuses to eat. The recipe at hand calls for one cup of broccoli to serve two people. How much broccoli needs to be prepared?
   1. 2 cups
   2. 3 cups
   3. 2.5 cups
   4. 3.5 cups

5. In a reaction, 5 moles reactant produces 3 moles of product. The following expression allows the calculation of the mole ratio reactant/product:
   1. \( \frac{3}{5} \)
   2. \( 5 - 3 \)
   3. \( \frac{2}{3} \)
   4. \( 5 + 3 \)

True/False:

6. _____ In question number one about the Super Bowl, if three additional people drop in at half-time, you will need to purchase one more bag of chips and one more 12-pack of soft drinks.
7. _____ The coefficients of a balance equation represent only the number of moles of materials.
8. _____ The production of ammonia from nitrogen and hydrogen is known as the Haber process.
9. _____ The number of moles of material is conserved when the reactants are converted to products.
10. _____ A mole ratio tells us directly the number of grams of reactant needed to produce product.

Short Answers:

11. $4FeS_2 + 11O_2 \rightarrow 2Fe_2O_3 + 8SO_2$
12. Calculate the mole ratio of
   1. FeS$_2$ to O$_2$
   2. FeS$_2$ to Fe$_2$O$_3$
   3. O$_2$ to SO$_2$
13. $3Ca + 2AlCl_3 \rightarrow 3CaCl_2 + 2Al$
14. Calculate the mole ratio of
   1. Ca to Al
   2. CaCl$_2$ to AlCl$_3$
   3. Ca to CaCl$_2$
15. $2NaOH + Cl_2 \rightarrow NaCl + NaClO + H_2O$
   1. How many moles of NaOH are needed to form 2.3 moles NaClO?
   2. How much NaCl will be formed from 1.7 moles Cl$_2$?
   3. How much NaCl will be formed from 3.1 moles NaOH?
16. $SiCl_4 + 4H_2O \rightarrow H_4SiO_4 + 4HCl$
   1. How much SiCl$_4$ is needed to produce 2.7 moles HCl?
   2. How many moles H$_4$SiO$_4$ will be formed from 3.9 moles SiCl$_4$?
   3. How much water is needed to produce 2.5 moles H$_4$SiO$_4$?
17. $CaC_2 + 2H_2O \rightarrow C_2H_2 + Ca(OH)_2$
   1. How many moles C$_2$H$_2$ can be formed from 8.3 moles CaC$_2$?
   2. How much water is required to produce 3.7 moles Ca(OH)$_2$?
   3. How many moles Ca(OH)$_2$ can be produced from 4.2 moles CaC$_2$?
18. $Al_4C_3 + 12H_2O \rightarrow 3CH_4 + 4Al(OH)_3$
   1. How many moles Al$_4$C$_3$ are needed to produce 4.6 moles CH$_4$?
   2. How many moles CH$_4$ are formed when the reaction produces 7.3 moles Al(OH)$_3$?
   3. How much Al(OH)$_3$ will be formed from 2.1 moles Al$_4$C$_3$?

Answer Key

1. d
2. c
3. b
4. c
5. c
6. true
7. false
8. true
9. false
10. false
12.1. Mole Ratios

1. \( \frac{4FeS_2}{11O_2} \)
2. \( \frac{4FeS_2}{2Fe_2O_3} = 2 : 1 \)
3. \( \frac{11O_2}{8SO_2} \)

1. \( \frac{3Ca}{2Al} \)
2. \( \frac{3CaCl_2}{2AlCl_3} \)
3. \( \frac{3Ca}{3CaCl_2} = 1 : 1 \)

1. \( 2.3 \) moles NaClO \( \times \frac{2NaOH}{NaClO} = 4.6 \) moles NaOH
2. \( 1.7 \) moles Cl\(_2\) \( \times \frac{NaCl}{Cl_2} = 1.7 \) mol NaCl
3. \( 3.1 \) moles NaOH \( \times \frac{NaCl}{NaOH} = 1.55 \) mol NaCl

1. \( 2.7 \) moles HCl \( \times \frac{SiCl_4}{HCl} = 0.68 \) moles SiCl\(_4\)
2. \( 3.9 \) moles SiCl\(_4\) \( \times \frac{H_4SiO_4}{SiCl_4} = 3.9 \) moles H\(_4\)SiO\(_4\)
3. \( 2.5 \) moles H\(_4\)SiO\(_4\) \( \times \frac{4H_2O}{H_4SiO_4} = 10.0 \) moles H\(_2\)O

1. \( 8.3 \) moles CaC\(_2\) \( \times \frac{C_2H_2}{CaC_2} = 8.3 \) moles C\(_2\)H\(_2\)
2. \( 3.7 \) moles Ca(OH)\(_2\) \( \times \frac{2H_2}{Ca(OH)_2} = 7.4 \) moles H\(_2\)O
3. \( 4.2 \) moles CaC\(_2\) \( \times \frac{Ca(OH)_2}{CaC_2} = 4.2 \) moles Ca(OH)\(_2\)

1. \( 4.6 \) moles CH\(_4\) \( \times \frac{Al_4C_3}{3CH_4} = 1.53 \) moles Al\(_4\)C\(_3\)
2. \( 7.3 \) moles Al(OH)\(_3\) \( \times \frac{3CH_4}{3Al(OH)_3} = 5.5 \) moles CH\(_4\)
3. \( 2.1 \) moles Al\(_4\)C\(_3\) \( \times \frac{4Al(OH)_3}{Al_4C_3} = 8.4 \) moles Al(OH)\(_3\)
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Short Answers:

1. \(\text{Ca}_3\text{P}_2 + 6\text{H}_2\text{O} \rightarrow 3\text{Ca} (\text{OH})_2 + 2\text{PH}_3\)
   1. How many moles of \(\text{Ca}_3\text{P}_2\) are needed to prepare 57.7 grams \(\text{PH}_3\)?
   2. How many moles of \(\text{Ca} (\text{OH})_2\) can be made from 100 grams \(\text{Ca}_3\text{P}_2\)?
   3. How many moles \(\text{PH}_3\) are formed in a reaction that yields 98.4 grams \(\text{Ca} (\text{OH})_2\)?

2. \(2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}\)
   1. How many liters of \(\text{CO}_2\) at STP are produced from the reaction of 4.7 moles \(\text{NaHCO}_3\)?
   2. How many grams of \(\text{NaHCO}_3\) are required to form 2.1 moles \(\text{Na}_2\text{CO}_3\)?
   3. How many grams carbon dioxide at STP are produced in a reaction that forms 1.9 moles sodium carbonate?

3. \(2\text{Au}_2\text{O}_3 \rightarrow 4\text{Au} + 3\text{O}_2\)
   1. What volume of \(\text{O}_2\) at STP is produced from the reaction of 212 grams of \(\text{Au}_2\text{O}_3\)?
   2. How many grams of \(\text{Au}\) are produced?
   3. How much \(\text{Au}_2\text{O}_3\) is needed to form 85 grams gold?

4. \(\text{Au}_2\text{S}_3 + 3\text{H}_2 \rightarrow 2\text{Au} + 3\text{H}_2\text{S}\)
   1. How many moles of \(\text{Au}_2\text{S}_3\) is required to form 56 grams of \(\text{H}_2\text{S}\) at STP?
   2. What volume of hydrogen at STP is needed to form 450 grams of gold?
   3. How many grams of \(\text{Au}_2\text{S}_3\) is needed to produce this amount of gold?

5. \(4\text{NH}_3 + 6\text{NO} \rightarrow 5\text{N}_2 + 6\text{H}_2\text{O}\)
   1. What volume of \(\text{N}_2\) at STP is produced from the reaction of 7.9 L \(\text{NO}\) with ammonia?
   2. How many liters of ammonia at STP will be consumed in the above process?
   3. If 2.4 moles of ammonia react with excess \(\text{NO}\), how many grams of \(\text{N}_2\) will be formed?

6. \(\text{FeCl}_3 + 3\text{NH}_4\text{OH} \rightarrow \text{Fe(OH)}_3 + 3\text{NH}_4\text{Cl}\)
   1. Excess ammonium hydroxide is reacted with 75.2 grams \(\text{FeCl}_3\). How many grams \(\text{Fe(OH)}_3\) are formed?
   2. How much ammonium hydroxide is needed to form 64 grams iron(III) hydroxide?
   3. How much iron(III) chloride is needed to react completely with 41 grams ammonium hydroxide?

Answer Key

1. \(57.7 \text{ g } \text{PH}_3 \times \frac{1 \text{ mol } \text{PH}_3}{34 \text{ g } \text{PH}_3} \times \frac{1 \text{ mol } \text{Ca}_3\text{P}_2}{2 \text{ mol } \text{PH}_3} = 0.85 \text{ moles } \text{Ca}_3\text{P}_2\)
2. \(100 \text{ g } \text{Ca}_3\text{P}_2 \times \frac{1 \text{ mol } \text{Ca}_3\text{P}_2}{182.3 \text{ g } \text{Ca}_3\text{P}_2} \times \frac{3 \text{ mol } \text{Ca(OH)}_2}{2 \text{ mol } \text{Ca}_3\text{P}_2} = 1.65 \text{ mol } \text{Ca(OH)}_2\)
3. \(98.4 \text{ g } \text{Ca(OH)}_2 \times \frac{1 \text{ mol } \text{Ca(OH)}_2}{75.1 \text{ g } \text{Ca(OH)}_2} \times \frac{2 \text{ PH}_3}{3 \text{ Ca(OH)}_2} = 0.87 \text{ mol } \text{PH}_3\)
1. \(4.7 \text{ mol NaHCO}_3 \times \frac{1 \text{ mol CO}_2}{2 \text{ mol NaHCO}_3} \times \frac{22.4 \text{ L}}{84 \text{ g mol}^{-1}} = 52.6 \text{ L CO}_2\)
2. \(2.1 \text{ mol Na}_2\text{CO}_3 \times \frac{2 \text{ mol NaHCO}_3}{1 \text{ mol Na}_2\text{CO}_3} \times \frac{44.0 \text{ g CO}_2}{\text{ mol}} = 352.8 \text{ g NaHCO}_3\)
3. \(1.9 \text{ mol Na}_2\text{CO}_3 \times \frac{1 \text{ mol CO}_2}{1 \text{ mol Na}_2\text{CO}_3} \times \frac{44.0 \text{ g CO}_2}{\text{ mol}} = 83.6 \text{ g CO}_2\)

1. \(212 \text{ g } \text{Au}_2\text{O}_3 \times \frac{\text{mol}}{442 \text{ g}} \times \frac{3 \text{ mol O}_2}{2 \text{ mol } \text{Au}_2\text{O}_3} \times \frac{22.4 \text{ L}}{\text{ mol}} = 16.1 \text{ L O}_2\)
2. \(212 \text{ g } \text{Au}_2\text{O}_3 \times \frac{\text{mol}}{442 \text{ g } \text{Au}_2\text{O}_3} \times \frac{4 \text{ mol Au}}{2 \text{ mol } \text{Au}_2\text{O}_3} \times \frac{197 \text{ g}}{\text{ mol Au}} = 189 \text{ g Au}\)
3. \(85 \text{ g Au} \times \frac{\text{mol}}{197 \text{ g}} \times \frac{2 \text{Au}_2\text{O}_3}{4 \text{Au}} \times \frac{442 \text{ g}}{\text{ mol } \text{Au}_2\text{O}_3} = 95.4 \text{ g } \text{Au}_2\text{O}_3\)

1. \(56 \text{ g } \text{H}_2\text{S} \times \frac{\text{mol}}{34.1 \text{ g}} \times \frac{1 \text{ mol } \text{Au}_2\text{S}_3}{3 \text{ mol } \text{H}_2\text{S}} = 0.55 \text{ mol } \text{Au}_2\text{S}_3\)
2. \(450 \text{ g Au} \times \frac{\text{mol}}{197 \text{ g}} \times \frac{3 \text{ mol H}_2}{2 \text{ mol Au}} \times \frac{22.4 \text{ L}}{\text{ mol}} = 76.8 \text{ L H}_2\)
3. \(450 \text{ g Au} \times \frac{\text{mol}}{197 \text{ g}} \times \frac{2 \text{ mol } \text{Au}_2\text{S}_3}{2 \text{ mol Au}} \times \frac{442 \text{ g}}{\text{ mol } \text{Au}_2\text{S}_3} = 504.8 \text{ g } \text{Au}_2\text{S}_3\)

1. \(7.9 \text{ L NO} \times \frac{5 \text{ L } \text{N}_2}{6 \text{ L NO}} = 5.83 \text{ L } \text{N}_2\)
2. \(7.9 \text{ L NO} \times \frac{4 \text{ L } \text{NH}_3}{6 \text{ L NO}} = 5.3 \text{ L } \text{NH}_3\)
3. \(2.4 \text{ mol } \text{NH}_3 \times \frac{5 \text{ mol N}_2}{4 \text{ mol } \text{NH}_3} \times \frac{28 \text{ g}}{\text{ mol } \text{N}_2} = 84.0 \text{ g } \text{N}_2\)

1. \(75.2 \text{ g } \text{FeCl}_3 \times \frac{\text{mol}}{162.4 \text{ g}} \times \frac{\text{mol } \text{Fe(OH)}_3}{\text{ mol } \text{FeCl}_3} \times \frac{106.9 \text{ g}}{\text{ mol } \text{Fe(OH)}_3} = 49.5 \text{ g } \text{Fe(OH)}_3\)
2. \(64 \text{ g } \text{Fe(OH)}_3 \times \frac{\text{mol}}{162.4 \text{ g}} \times \frac{3 \text{ mol } \text{NH}_4\text{OH}}{\text{ mol } \text{Fe(OH)}_3} \times \frac{35.1 \text{ g}}{\text{ mol } \text{NH}_4\text{OH}} = 63.0 \text{ g } \text{NH}_4\text{OH}\)
3. \(41 \text{ g } \text{NH}_4\text{OH} \times \frac{\text{mol}}{35.1 \text{ g}} \times \frac{3 \text{ mol } \text{FeCl}_3}{\text{ mol } \text{NH}_4\text{OH}} \times \frac{162.4 \text{ g}}{\text{ mol } \text{FeCl}_3} = 63.2 \text{ g } \text{FeCl}_3\)
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. After a reaction was run, there were 12.7 grams of reagent B left over. Reagent B is the
   1. superfluous reagent
   2. excess reagent
   3. extra reagent
   4. limiting reagent

2. Reagent A was used in a reaction. At the end of the reaction, no reagent A remained. Reagent A is the
   1. excess reagent
   2. insufficient reagent
   3. limiting reagent
   4. restricted reagent

3. The amount of product that is expected to be formed is the
   1. theoretical yield
   2. predicted yield
   3. limiting yield
   4. desired yield

4. The amount of product seen at the end of a chemical reaction is used to determine
   1. desired yield
   2. actual yield
   3. measured yield
   4. predicted yield

5. In order to calculate percent yield, all of the following information is needed except
   1. theoretical yield
   2. amount actually obtained
   3. expected amount
   4. amount of excess reagent

True/False:

6. _____ The presence of moisture in the product does not affect determination of actual yield.
7. _____ The product must be pure in order to determine actual yield.
8. _____ The excess reagent is always due to incomplete reactions.
9. _____ Excess reagent is calculated from mole data.
10. _____ A balanced equation is necessary for theoretical yield calculations.
12.3 Limiting Reactant and Percent Yield

Short Answers:

**Part A – Limiting Reagent and Excess Reagent**

For problems 11-15, determine

a. the limiting reagent
b. the excess reagent
c. the number of moles of excess reagent

11. 25.0 grams of chlorine reacts with 29.0 grams of sodium
12. \( 2Na + Cl_2 \rightarrow 2NaCl \)
13. 2.80 g Mg reacts with 11.0 g \( O_2 \)
14. \( 2Mg + O_2 \rightarrow 2MgO \)
15. 75.4 grams of \( C_2H_3Br_3 \) reacts with 47.1 grams of \( O_2 \)
16. \( 4C_2H_3Br_3 + 11O_2 \rightarrow 8CO_2 + 6H_2O + 6Br_2 \)
17. 22.5 grams of \( CoO \) reacts with 2.6 grams of \( O_2 \)
18. \( 4CoO + O_2 \rightarrow 2Co_2O_3 \)
19. 21 grams sodium nitrate reacts with 15 grams sulfuric acid
20. \( 2NaNO_3(s) + H_2SO_4(l) \rightarrow Na_2SO_4(s) + 2HNO_3(g) \)

**Part B – Theoretical Yield, Percent Yield**

For each of the following reactions, calculate

a. theoretical yield
b. percent yield

16. 65 grams of \( CaCO_3 \) is heated to give 18 grams of \( CaO \)
17. \( CaCO_3 \rightarrow CaO + CO_2 \)
18. 59 mol of \( Ca_3(PO_4)_2 \) is used and 95 mol of \( CaSiO_3 \) is obtained
19. \( Ca_3(PO_4)_2 + 3SiO_2 + 5C \rightarrow 3CaSiO_3 + 2P + 5CO \)
20. 55.3 g \( WO_3 \) yields 40.7 g of tungsten
21. \( WO_3 + 3H_2 \rightarrow W + 3H_2O \)
22. 420 g \( CS_2 \) produces 740 g \( CCl_4 \)
23. \( CS_2 + 3Cl_2 \rightarrow CCl_4 + S_2Cl_2 \)
24. 0.40 g \( NO_2 \) forms 0.39 g \( N_2O_5 \)
25. \( 2NO_2 + O_3 \rightarrow N_2O_5 + O_2 \)

**Answer Key**

1. b
2. c
3. a
4. b
5. d
6. false
7. true
8. false
9. true
10. true
11. 29 g Na $\times \frac{mol}{23 g} = 1.26 mol Na$  
25 g $Cl_2 \times \frac{mol}{71 g} = 0.35 mol Cl_2$
   1. $Cl_2$ is the limiting reagent
   2. Na is the excess reagent.
   3. for 0.35 mol $Cl_2$, need 0.7 mol Na, so 0.56 mol Na in excess

12. 2.8 g Mg $\times \frac{mol}{24.3 g} = 0.12 mol Mg$  
11.0 g $O_2 \times \frac{mol}{32 g} = 0.34 mol O_2$
   1. Mg is the limiting reagent
   2. $O_2$ is the excess reagent
   3. for 0.12 mol Mg, need 0.06 mol $O_2$, so 0.28 mol $O_2$ in excess

13. 0.12 mol Mg available, need 0.68 mol to react with $O$
   1. $C_2H_3Br_3$
   2. $C 2 \times 12.0 = 24.0$
   3. $H 3 \times 1.01 = 3.03$
   4. $Br 3 \times 79.9 = \underline{239.7}$
      266.7 g/mol

17. 75.4 g $\times \frac{mol}{266.7 g} = 0.28 mol C_2H_3Br_3$
18. 47.1 g $O_2 \times \frac{mol}{32.0 g} = 1.47 mol O_2$
19. 0.28 mol $C_2H_3Br_3 \times 4 \frac{mol O_2}{mol C_2H_3Br_3} = 0.77 mol O_2$
   1. $C_2H_3Br_3$ is the limiting reagent
   2. $O_2$ is the excess reagent
   3. 0.7 mol $O_2$ in excess

20. 22.5 g $CoO \times \frac{mol}{74.9 g} = 0.30 mol CoO$
21. 2.6 g $O_2 \times \frac{mol}{32.0 g} = 0.08 mol$
22. 0.30 mol $CoO \times 4 \frac{mol O_2}{mol CoO} = 0.075 mol O_2$
   1. CoO is the limiting reagent
   2. $O_2$ is the excess reagent
   3. 0.005 mol $O_2$ in excess

23. NaNO$_3$
24.
   $Na 1 \times 23 = 23$
   $N 1 \times 14 = 14$
   $O 3 \times 16 = 48$
   $85 g/mol NaNO_3$

25. H
26. 2
27. SO
28. 4
29.
   $H 2 \times 1.01 = 2.02$
   $S 1 \times 32.1 = 32.1$
   $O 4 \times 16 = 64$
   $98.1 g/mol H_2SO_4$
30. $21 \text{ g } \text{NaNO}_3 \times \frac{\text{mol}}{85 \text{ g}} = 0.25 \text{ mol } \text{NaNO}_3$

31. $15 \text{ g } \text{H}_2\text{SO}_4 \times \frac{\text{mol}}{98.1 \text{ g}} = 0.15 \text{ mol } \text{H}_2\text{SO}_4$

32. $0.25 \text{ mol } \text{NaNO}_3 \times \frac{\text{mol } \text{H}_2\text{SO}_4}{\sum \text{mol } \text{NaNO}_3} = 0.125 \text{ mol } \text{H}_2\text{SO}_4$
   
   1. NaNO$_3$ is the limiting reagent
   2. H$_2$SO$_4$ is the excess reagent
   3. 0.025 mol H$_2$SO$_4$ in excess

33. CaCO$_3$

$$Ca \times 40.1 = 40.1$$
$$C \times 12.0 = 12.0$$
$$O \times 16.0 = 48$$

100.1 g/mol

35. CaO

$$Ca \times 40.1 = 40.1$$
$$O \times 16.0 = 16.0$$

56.1 g/mol

37. $65 \text{ g } \text{CaCO}_3 \times \frac{\text{mol}}{100.1 \text{ g}} \times \frac{\text{mol } \text{CaO}}{\text{mol } \text{CaCO}_3} \times \frac{56.1 \text{ g}}{\text{mol } \text{CaO}} = 36.4 \text{ g } \text{CaO} \text{ theoretical}$

38. percent yield $= \frac{18.6 \text{ g}}{36.4 \text{ g}} \times 100\% = 51.5\%$

39. $59 \text{ mol } \text{Ca}_3(\text{PO}_4)_2 \times \frac{3 \text{ mol } \text{CaSiO}_3}{\text{mol } \text{Ca}_3(\text{PO}_4)_2} = 177 \text{ mol } \text{CaSiO}_3 \text{ theoretical yield}$

40. percent yield $= \frac{95 \text{ mol}}{177 \text{ mol}} \times 100\% = 53.7\%$

41. WO$_3$

$$W = 1 \times 183.8 = 183.8$$
$$O = 3 \times 16.0 = 48$$

231.8 g/mol

43. $55.3 \text{ g } \text{WO}_3 \times \frac{\text{mol}}{231.8 \text{ g}} \times \frac{\text{mol } \text{W}}{\text{mol } \text{WO}_3} \times \frac{183.8 \text{ g}}{\text{mol } \text{W}} = 43.8 \text{ g } \text{W}$

44. percent yield $= \frac{40.7 \text{ g}}{43.8 \text{ g}} \times 100\% = 92.9\%$ yield

45. CS$_2$

$$C \times 12.0 = 12.0$$
$$S \times 32.1 = 64.2$$

76.2 g/mol

47. CCl

48. 4

49. $C \times 12.0 = 12.0$
$$\text{Cl} \times 35.5 = 142$$

154 g/mol
50. \[420 \text{ g } \text{CS}_2 \times \frac{\text{mol}}{76.2 \text{ g}} \times \frac{\text{mol } \text{CCl}_4}{\text{mol } \text{CS}_2} \times \frac{154 \text{ g}}{\text{mol}} = 848.8 \text{ g } \text{CCl}_4 \text{ theoretical}\]

51. \[\text{percent yield} = \frac{740 \text{ g}}{848.8 \text{ g}} \times 100\% = 87.2\%\]

52. \[\text{NO}_2 = 46 \text{ g/mol } \text{N}_2\text{O}_5 = 108 \text{ g/mol}\]

53. \[0.40 \text{ g } \text{NO}_2 \times \frac{\text{mol } \text{N}_2\text{O}_5}{2 \text{ mol } \text{NO}_2} \times \frac{108 \text{ g}}{\text{mol } \text{N}_2\text{O}_5} = 0.47 \text{ g } \text{N}_2\text{O}_5\]

54. \[\text{percent yield} = \frac{0.39 \text{ g}}{0.47 \text{ g}} \times 100\% = 83\%\]
Lesson Quiz

1. A certain reaction produces 56.2 grams of product. This data is called the
   1. theoretical yield
   2. percent yield
   3. actual yield
   4. predicted yield

2. At the completion of a reaction, 6.3 grams of reagent A were left over. Reagent A is the
   1. excess reagent
   2. extra reagent
   3. limiting reagent
   4. remaining reagent

3. You are preparing pizzas to eat while watching the NCAA championship game on TV. Your recipe calls for
   27 slices of pepperoni for each pizza. You have 142 slices of pepperoni. How many complete pizzas can you
   prepare?
   1. four
   2. six
   3. five
   4. seven

4. In the reaction $A + B \rightarrow C + D$, 7.2 moles A form 5.4 moles C. The mole ratio A:C is
   1. 2:1
   2. 4:2
   3. 3:2
   4. 4:3

5. The coefficients in a chemical equation represent the ________ of the individual materials
   1. number of grams
   2. number of moles
   3. volume
   4. molar designations

6. In the reaction for the Haber process, $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$. The mole ratio of hydrogen to ammonia
   can be expressed in any of the following ways except
   1. 3:2
   2. 9:6
   3. 5:3
   4. 6:4
7. In the reaction \( C_3H_8 + 5O_2 \rightarrow 4H_2O + 3CO_2 \), you react 2.5 moles \( C_3H_8 \) with 14.7 moles oxygen. The reaction goes to completion with 100% yield. What reagent is left over and how much of that reagent remains:
   1. 0.23 moles \( C_3H_8 \)
   2. 2.2 moles \( O_2 \)
   3. 1.7 moles \( C_3H_8 \)
   4. 1.6 mole \( O_2 \)

8. The amount of a given substance in moles x mole ratio = ?
   1. moles of unknown substance
   2. percent yield
   3. grams of unknown substance
   4. theoretical yield

True/False:
9. _____ The limiting reagent is the amount of material remaining after the reaction is completed.
10. _____ A balanced equation can be used to determine actual yield of a reaction.
11. _____ Actual yield is not affected by purity of product.
12. _____ Mass must be conserved in a chemical reaction.
13. _____ The mole ratio allows conversion from moles of given substance to moles of unknown substance.
14. _____ Side reactions do not influence the actual yield.
15. _____ Percent yield is obtained by dividing actual yield by theoretical yield and multiplying by 100%.
16. _____ In the laboratory, amounts of materials are usually expressed in terms of mass.

Short Answers:
17. 1. In the reaction \( 3Br_2 + 2FeI_3 \rightarrow 3I_2 + 2FeBr_3 \), how many moles of iodine are produced if 4.7 moles bromine are reacted with excess iron(III) iodide?
2. How many moles of \( FeI_3 \) are needed to react completely with 145 g \( Br_2 \)?
3. What is the theoretical yield of \( I_2 \) under the conditions of problem (c) above?
4. The reaction produced 197.2 g \( I_2 \). What is the percent yield?

1. In the reaction \( I_2O_5(g) + 5CO(g) \rightarrow 5CO_2(g) + I_2(g) \), how many moles of carbon monoxide is needed to react with 47.2 L of \( I_2O_5 \) at STP?
2. How many liters of \( I_2 \) are formed under the above conditions?
3. The reaction is carried out and 469.7 g \( I_2 \) are formed. What is the percent yield?

18. Use the reaction \( 3Zn + 2MoO_3 \rightarrow Mo_2O_3 + 3ZnO \) to solve the following problems.
1. What is the limiting reagent when 20.0 g \( MoO_3 \) reacts with 10.0 g \( Zn \)?
2. How much \( ZnO \) will be formed under these conditions?
3. If the percent yield for the reaction is 72%, how much \( ZnO \) was actually formed?

19. Use the reaction \( F_2(g) + MgCl_2(s) \rightarrow Cl_2(g) + MgF_2(s) \) to answer the following questions:
1. What volume of \( Cl_2 \) gas is formed at STP from the reaction of 7.2 mol \( F_2 \)?
2. How many liters of \( Cl_2 \) are formed at STP from 24.6 g \( MgCl_2 \)?
3. How many grams of \( F_2 \) are needed to carry out the reaction in (b)?

Answer Key
1. c
2. a  
3. c  
4. d  
5. b  
6. c  
7. b  
8. a  
9. false  
10. false  
11. false  
12. true  
13. true  
14. false  
15. true  
16. true

1. \[ \text{4.7 moles Br}_2 \times \frac{3 \text{ moles I}_2}{3 \text{ moles Br}_2} = 4.7 \text{ moles I}_2 \]

2. \[ \text{145 g Br}_2 \times \frac{\text{mol}}{159.8 \text{ g}} \times \frac{2 \text{ mol FeI}_3}{3 \text{ mol Br}_2} = 0.60 \text{ mol FeI}_3 \]

3. \[ \text{145 g Br}_2 \times \frac{\text{mol}}{159.8 \text{ g}} \times \frac{3 \text{ mol I}_2}{3 \text{ mol Br}_2} \times \frac{253.8 \text{ g}}{\text{mol I}_2} = 230.3 \text{ g I}_2 \text{ theoretical yield} \]

4. \[ \text{197.2 g} \times 100\% = 85.6\% \text{ yield} \]

1. \[ \text{47.2 L I}_2\text{O}_5 \times \frac{\text{mol}}{22.4 \text{ L}} \times \frac{5 \text{ mol CO}}{5 \text{ mol I}_2\text{O}_5} = 10.5 \text{ mol CO} \]

2. \[ 10.5 \text{ mol CO} \times \frac{\text{mol I}_2}{5 \text{ mol CO}} \times \frac{253.8 \text{ g}}{\text{mol I}_2} = 533 \text{ g I}_2 \text{ formed} \]

3. \[ \frac{469.7 \text{ g actual}}{533 \text{ g theoretical}} \times 100\% = 88.1\% \text{ yield} \]
Chapter 13

States of Matter Assessments

Chapter Outline

13.1 Kinetic - Molecular Theory and Gases
13.2 Liquids
13.3 Solids
13.4 Changes of State
13.5 States of Matter
Lesson Quiz

Multiple Choice:

1. One of the following is not a part of the kinetic-molecular theory
   1. matter is composed of tiny particles
   2. particles are in constant motion
   3. only atoms are affected by the theory
   4. kinetic energy is related to temperature

2. Ideal gas particles are all of the following except
   1. small spheres
   2. close together
   3. volume of particles insignificant
   4. in constant motion

3. Gas particles are
   1. strongly attracted toward one another
   2. connected by dipole-dipole forces
   3. have no intermolecular attractions
   4. connected by van der Waals forces

4. Gas exert pressure because of
   1. the mass of the particles
   2. the intermolecular forces
   3. the closeness of the particles
   4. the rapid motions of the particles

5. A mercury barometer measures
   1. atmospheric pressure
   2. tire pressure
   3. pressure inside a container
   4. vapor pressure

6. One of the following statements about atmospheric pressure is not true
   1. pressure due to downward force of air particles
   2. pressure due to upward motion of air particles
   3. higher at lower altitudes
   4. can be related to mm Hg

7. Absolute zero is the temperature at which
   1. hydrogen freezes
   2. the Kelvin scale no longer works
3. there is no molecular motion
4. all atomic vibrations are at a maximum

8. The Kelvin temperature is an indication of
   1. molecular stretching
   2. kinetic energy of the particles
   3. reactivity of the molecules
   4. size of the molecules

9. Pressure can be calculated using the formula
   1. force ÷ area
   2. mass ÷ acceleration
   3. area ÷ force
   4. area ÷ mass

10. Vapor pressure is determined using a
    1. manometer
    2. mercury barometer
    3. aneroid barometer
    4. water barometer

**True/False:**

11. _____ The barometer was invented by Luigi Torricelli
12. _____ One standard atmospheric pressure is equivalent to 760 torr.
13. _____ Gravity is responsible for gas pressure inside a balloon.
14. _____ Collisions between gas particles are elastic collisions.
15. _____ The kinetic-molecular theory applies to all particles.
16. _____ Real gases behave exactly like ideal gases.
17. _____ Molecules of H₂ in the gas phase interact by way of hydrogen bonds.
18. _____ A given mass exerts more pressure on a large surface area than on a small surface area.
19. _____ Gravity holds atmospheric gases in place.
20. _____ Kinetic energy increases as the temperature increases.

**Fill in the Blank:**

21. List the five points of the kinetic-molecular theory.
22. Define the following terms:
   1. pressure
   2. ideal gas
   3. atmospheric pressure
   4. absolute zero

23. Make the following conversions:
   1. 1.7 atm to torr
   2. 780 mm Hg to kPa
   3. 245 kPa to atm
   4. 450 kPa to mm Hg
   5. 987 torr to atm
1. Gases consist of very large numbers of tiny spherical particles that are far apart from one another compared to their size.
2. Gas particles are in constant rapid motion in random directions.
3. Collisions between gas particles and between particles and the container walls are elastic collisions.
4. There are no forces of attraction or repulsion between gas particles.
5. There are no forces of attraction or repulsion between gas particles.

1. pressure - the force per unit area on a surface.
2. ideal gas - an imaginary gas whose behavior perfectly fits all the assumptions of the kinetic-molecular theory.
3. atmospheric pressure - the pressure exerted by the gas particles in Earth’s atmosphere as those particles collide with objects.
4. absolute zero - the temperature at which the motion of particles theoretically ceases.

1. $1.7 \text{ atm} \times \frac{760 \text{ torr}}{\text{atm}} = 1292 \text{ torr}$
2. $780 \text{ mm Hg} \times \frac{101.3 \text{ kPa}}{760 \text{ mm Hg}} = 104.0 \text{ kPa}$
3. $245 \text{ kPa} \times \frac{1\text{ atm}}{101.3 \text{ kPa}} = 2.42 \text{ atm}$
4. $450 \text{ kPa} \times \frac{760 \text{ mm Hg}}{101.3 \text{ kPa}} = 3376 \text{ mm Hg}$
5. $987 \text{ torr} \times \frac{\text{atm}}{760 \text{ torr}} = 1.3 \text{ atm}$
Lesson Quiz

Name___________________ Class______________ Date ________

Multiple Choice:

1. In liquids, the intermolecular forces are
   1. weak
   2. intermittent
   3. strong
   4. minor

2. One similarity between liquids and gases
   1. both types of particles are close together
   2. both are fluids
   3. both have definite shapes
   4. both types of particles are far apart

3. Surface tension is a measure of
   1. elastic force in the liquid’s surface
   2. pressure against the sides of the container
   3. resistance to flow
   4. ability to float solid objects

4. One of the following statements about evaporation of a liquid is true
   1. only occurs above the boiling point of a liquid
   2. occurs at temperatures below the boiling point of a liquid
   3. higher temperatures decrease the rate of evaporation
   4. takes place only in closed containers

5. One of the following statements about liquid in a closed container is not true
   1. particles in the vapor state collide against the walls of the container
   2. liquids with weak intermolecular forces evaporate more quickly
   3. particles with strong intermolecular forces have high vapor pressures
   4. vapor pressure value depends upon the specific liquid

6. The standard pressure for measuring boiling point is
   1. 1.2 atm
   2. 105.9 kPa
   3. 760 torr
   4. 740 mm Hg

7. The boiling point of ethanol will be lowest
   1. on a mountain top
   2. at the sea shore
8. The intermolecular interaction that will produce the highest surface tension is
   1. dipole-dipole interactions
   2. ionic bonds
   3. hydrogen bonds
   4. van der Waals forces

9. Vapor pressure
   1. can be measured with a barometer
   2. increases with an increase in temperature
   3. is higher with particles that have strong intermolecular interactions
   4. is not affected by molecular structure

10. Boiling point
    1. is lower at higher barometric pressure
    2. is higher at low barometric pressure
    3. is lower at low barometric pressure
    4. is not affected by barometric pressure

True/False:

11. _____ In a closed container, the rate of evaporation is greater than the rate of condensation.
12. _____ Liquids with particles of higher kinetic energy will evaporate more quickly than a liquid whose particles have low kinetic energy.
13. _____ Diethyl ether has a higher vapor pressure than water.
14. _____ At a given temperature, all the particles of a liquid have the same kinetic energy.
15. _____ Liquids are less dense than gases.
16. _____ Liquids are classified as fluids.
17. _____ A rock will float on water because of the high surface tension of the liquid.
18. _____ At the boiling point of a liquid, all the particles have enough kinetic energy to evaporate.
19. _____ A manometer uses a U-tube to determine vapor pressure.
20. _____ A dynamic equilibrium exists when the rate of evaporation is greater than the rate of condensation.

Fill in the Blank:

21. Define the following terms:
    1. fluid
    2. surface tension
    3. vaporization
    4. evaporation
    5. condensation

22. __________ is a measure of the pressure exerted by the vapor that forms _____ its liquid form in a sealed container.

23. When the liquid in a closed container is _____, more molecules escape the liquid phase and evaporate.

24. Which of the following molecules will have the lowest vapor pressure?
    1. CH₃CH₂CH₃
    2. CH₃CH₂NH₂
    3. CH₃CH₂-O-CH₂CH₃

25. In order for a molecule to escape into the ____ state, it must have enough ________ energy to overcome the ______________ attractive forces in the liquid.
26. What is the difference between vaporization and evaporation?

---

**Answer Key**

1. c
2. b
3. a
4. b
5. c
6. c
7. a
8. c
9. b
10. c
11. false
12. true
13. true
14. false
15. false
16. true
17. false
18. true
19. true
20. false

1. fluid - a substance that is capable of flowing from one place to another and takes the shape of its container.
2. surface tension - a measure of the elastic force in the liquid’s surface.
3. vaporization - the process in which a liquid is converted to a gas.
4. evaporation - the conversion of a liquid to its vapor form below the boiling temperature of the liquid.
5. condensation - the change of state from a gas to a liquid.

21. vapor pressure, above
22. heated
23. molecule (b) –it can form hydrogen bonds
24. gas, kinetic, intermolecular
25. Vaporization is an overall term that simply refers to a liquid being converted to a gas. Evaporation is a specific type of vaporization that depends on the temperature.
Lesson Quiz

Name___________________ Class_________________ Date ________

Multiple Choice:

1. One of the following is not true of solids
   1. particles are closely packed
   2. very compressible
   3. solid form most dense of all three forms of matter
   4. a condensed state

2. Solid carbon dioxide
   1. melts readily
   2. converts to a liquid at -78°C
   3. converts directly to a gas
   4. is a solid at room temperature

3. Solids can be classified into ____ crystal systems.
   1. three
   2. seven
   3. four
   4. six

4. One of the following is not a crystal system
   1. biclinic
   2. monoclinic
   3. triclinic
   4. cubic

5. In the tetragonal crystal system
   1. $a \neq b \neq c; \alpha = \beta = \gamma = 90^\circ$
   2. $a = b = c; \alpha = \beta = \gamma = 90^\circ$
   3. $a \neq b \neq c; \alpha \neq \beta \neq \gamma \neq 90^\circ$
   4. $a = b \neq c; \alpha = \beta = \gamma = 90^\circ$

6. In the simple cubic system unit cell, atoms or ions are
   1. at the corners and multiple atoms or ions are in the center of each of the six faces
   2. at the corners and a single atom or ion in the center of each of the six faces
   3. only at the corners of the unit cell
   4. only in the center of the unit cell

7. One of the following statements about ionic crystals is not true
   1. generally formed between metals and group 16 or 17 non-metals
   2. brittle
3. some contain polyatomic ions
4. low melting points

8. One of the following is not true about metallic crystals:
   1. consist of cations
   2. have very mobile electrons
   3. narrow range of melting points
   4. good conductors of electricity

9. Covalent network crystals
   1. are connected by ionic bonds
   2. are hard and brittle
   3. conduct electricity well in molten state
   4. have low melting points

10. One of the following is not an amorphous solid
    1. diamond
    2. glass
    3. plastic
    4. rubber

True/False:

11. _____ Nonmetallic polyatomic ions are often found in ionic crystals.
12. _____ The edge lengths of a crystal are represented by $\alpha$, $\beta$, and $\gamma$.
13. _____ Solids have a vapor pressure.
14. _____ The melting point of solid aluminum is the same as the freezing point of liquid aluminum.
15. _____ The hexagonal crystal structure can be represented by $a = b = c; \alpha = \beta = \gamma \neq 90^\circ$.
16. _____ The majority of solids are crystalline in nature.
17. _____ The cubic crystal system is composed of four different types of unit cells.
18. _____ Boron is an example of a covalent network crystalline solid.
19. _____ Molecular crystals are connected by covalent bonds.
20. _____ Glass can be considered to be a supercooled liquid.

Fill in the Blank:

21. Identify the following unit cells:

   ![Unit Cell Diagram](image.png)
22. Define the following terms:
   1. crystal
   2. sublimation
   3. deposition
   4. melting point
   5. unit cell

23. Match the cell dimensions with the appropriate cell:
   1. \( a = b = c; \alpha = \beta = \gamma = 90^\circ \quad 1. \) tetragonal
   2. \( a = b \neq c; \alpha = \beta = \gamma = 90^\circ \quad 2. \) cubic
   3. \( a \neq b \neq c; \alpha = \beta = \gamma = 90^\circ \quad 3. \) orthorhombic

24. List two examples of
   1. ionic crystals
   2. metallic crystals
   3. covalent network crystals
   4. molecular crystals

---

**Answer Key**

1. b
2. c
3. b
4. a
5. d
6. c
7. d
8. c
9. d
10. a
11. true
12. false
13. true
14. true
15. false
16. true
17. false
18. true
19. false
20. true

1. body-centered cubic
2. face-centered cubic
3. simple cubic

1. crystal - a substance in which the particles are arranged in an orderly, repeating, three-dimensional pattern.
2. sublimation - the change of state from a solid to a gas without passing through the liquid state.
3. deposition - the change of state from a gas to a solid.
4. melting point - the temperature at which a solid changes into a liquid.
5. unit cell - the smallest portion of a crystal lattice that shows the three-dimensional pattern of the entire crystal.

1. cubic
2. tetragonal
3. orthorhombic

1. NaCl, CaF₂
2. Hg, Na, Au, W
3. B, C(diamond), SiO₂
4. H₂, I₂, NH₃, H₂O
Lesson Quiz

Name___________________ Class______________ Date ________

Multiple Choice:

1. The hydrogen bonds in frozen water begin to break apart
   1. as the ice warms
   2. at 0°C
   3. at 0°C
   4. after the ice has completely melted

2. During the melting process
   1. the temperature of the system gradually increases
   2. the temperature remains constant
   3. the temperature of the liquid portion increases
   4. the temperature of the ice increases

3. Water vaporizes to steam at a temperature of
   1. 100 K
   2. 273 K
   3. 373 K
   4. 150 K

4. The transition from a liquid to a solid is known as
   1. freezing
   2. condensation
   3. sublimation
   4. vaporization

5. The change of state behavior for any material can be represented by a
   1. temperature curve
   2. condensing curve
   3. heating curve
   4. melting curve

6. A phase diagram for a material provides all the following information except
   1. the behavior of the gas under pressure
   2. the behavior of the solid under pressure
   3. the effect of pressure on the liquid to gas transition
   4. the effect of impurities in the material

7. All of the following statements about the water phase diagram are true except
   1. an increase in pressure increases the boiling point
   2. a decrease in pressure lowers the freezing point
3. a decrease in temperature lowers the pressure needed for boiling
4. an increase in temperature at constant pressure facilitates the shift from liquid to gas

8. Steam condenses to water at a temperature of
   1. 273 K
   2. 473 K
   3. 100°C
   4. 0°C

9. The equilibrium between the solid and gas states is represented by
   1. freezing
   2. condensation
   3. deposition
   4. vaporization

10. A cooling curve represents all of the following except
    1. the transition from vapor to liquid
    2. the transition from liquid to solid
    3. the overall change from vapor to solid
    4. the transition from solid to vapor

True/False:

11. _____ Temperature is always constant during a change of state
12. _____ Phase diagrams only show solid to liquid transitions.
13. _____ The heating curve for carbon dioxide at one atm pressure shows two plateaus.
14. _____ Melting of ice involves disruption of hydrogen bonds.
15. _____ The cooling curve is the reverse of the heating curve.
16. _____ Ice is less dense than liquid water.
17. _____ Ice at -5°C can be converted to water by decreasing the pressure.
18. _____ Under the right conditions, solid ice can be directly changed to vapor.
19. _____ Melting and boiling points cannot be determined from a heating curve.
20. _____ Hydrogen bonds are important in both the solid and liquid stages of water.

Fill in the Blank:

21. Define the following terms:
    1. freezing
    2. deposition
    3. triple point
    4. critical pressure
    5. phase diagram

22. Fill in the blanks:
    1. At a sufficiently low pressure, the ________ phase does not exist.
    2. As _____ is steadily added to the ice block, the water molecules will begin to ________ faster and faster
       as they absorb __________ energy.

23. Use the phase diagram for carbon dioxide below to answer the following questions:
13.4. Changes of State

1. At 250 K and 1 bar pressure, what is the physical state of CO$_2$?
2. At 100 bar pressure and 250 K, what is the physical state of CO$_2$?
3. At one bar pressure and 250 K temperature, CO$_2$ is a gas. How can it be sold as dry ice?

**Answer Key**

1. c
2. b
3. c
4. a
5. c
6. d
7. b
8. c
9. c
10. d
11. true
12. false
13. false
14. true
15. true
16. true
17. false
18. true
19. false
20. true

1. freezing – transition from liquid to solid
2. deposition – transition from vapor to solid
3. triple point - the only temperature/pressure pairing at which the solid, liquid, and vapor states of a substance can all coexist at equilibrium.
4. critical pressure ($P_c$) - the pressure that must be applied to the gas at the critical temperature in order to turn it into a liquid.
5. phase diagram - a graph showing the conditions of temperature and pressure under which a substance exists in the solid, liquid, and gas phases.

1. liquid
2. heat, vibrate, kinetic

1. gas
2. liquid
3. The storage temperature is kept low enough that the material remains in the solid state.
Chapter Test

Multiple Choice:

1. One of the following is not a form of matter
   1. solid
   2. liquid
   3. vacuum
   4. vapor

2. Pressure equals
   1. \( \frac{\text{force}}{\text{area}} \)
   2. \( \frac{\text{area}}{\text{force}} \)
   3. \( \text{force} \times \text{area} \)
   4. \( \text{force} + \text{area} \)

3. Atmospheric pressure is measured with a
   1. manometer
   2. barometer
   3. telemeter
   4. sphygmomanometer

4. The kinetic energy in a sample of sulfur
   1. is increased with a decrease in temperature
   2. decreases as the temperature goes up
   3. is not affected by temperature
   4. increases as the temperature increases

5. Gases and liquids have one property in common.
   1. both are compressible
   2. both are fluids
   3. both have particles that are close together
   4. both have strong interactions among the particles of the material

6. The elastic force in the surface of a liquid is called
   1. surface tension
   2. surface pressure
   3. surface attraction
   4. surface energy

7. Vapor pressure is influenced by all of the following except
   1. temperature
   2. atmospheric pressure
   3. molecular mass
   4. intermolecular forces
8. A manometer is used to measure
   1. atmospheric pressure
   2. pressure in a closed reaction vessel
   3. reaction pressure
   4. vapor pressure
9. Solids and liquids have the following property in common
   1. particles close together
   2. both flow freely
   3. both have defined shape
   4. both take the shape of the container
10. Metallic crystals
    1. are poor conductors of electricity
    2. consist of anions
    3. have mobile electrons
    4. have a narrow range of melting points

True/False:

11. _____ The kinetic-molecular theory applies to all forms of matter.
12. _____ When gas particles collide, kinetic energy is increased.
13. _____ A heavier rock exerts the same pressure on the ground as a lighter rock of the same dimensions.
14. _____ Intermolecular forces hold liquid particles together.
15. _____ High kinetic energy in a liquid decreases the evaporation rate.
16. _____ Evaporation of a liquid occurs below the boiling point.
17. _____ A barometer is a system for measuring atmospheric pressure.
18. _____ Solid particles are organized in regular three-dimensional arrays.
19. _____ Surface tension measures the pull of liquid particles by the walls of the containing vessel.
20. _____ The edge lengths of a crystal are presented by a, b, and c.

Fill in the Blank:

21. Define the following terms:
   1. pressure
   2. evaporation
   3. unit cell
   4. critical pressure
22. Make the following conversions:
   1. 972 torr to atm
   2. 302 kPa to torr
   3. 4.7 atm to mm Hg
   4. 845 mm Hg to kPa
23. Which of the molecules below will demonstrate the lowest surface tension?
   1. \( CH_3CH_2CH_2 – OH \)
   2. \( CH_3CH – O – CH_3 \)
      \[ \| \]
      \( NH_2 \)
3. \(CH_2 = CH_2\)

24. Use the graph below to answer the following questions about chloroform:

1. What is the approximate vapor pressure at 60°C?
2. At what temperature is the vapor pressure 10,000 mm Hg?

26. Use the phase diagram for water to answer the following questions:

1. If the pressure on the system is increased, what happens to the temperature needed to freeze water?
2. How does an increase in pressure influence the boiling point of water?
6. a  
7. c  
8. b  
9. a  
10. c  
11. true  
12. false  
13. false  
14. true  
15. false  
16. true  
17. true  
18. true  
19. false  
20. true

1. pressure - the force per unit area on a surface.
2. evaporation - the conversion of a liquid to its vapor form below the boiling temperature of the liquid.
3. unit cell - the smallest portion of a crystal lattice that shows the three-dimensional pattern of the entire crystal.
4. critical pressure - the pressure that must be applied to the gas at the critical temperature in order to turn it into a liquid.

1. \(972 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 1.28 \text{ atm}\)
2. \(302 \text{ kPa} \times \frac{1 \text{ atm}}{101.3 \text{ kPa}} = 2266 \text{ torr}\)
3. \(4.7 \text{ atm} \times \frac{760 \text{ mm Hg}}{1 \text{ atm}} = 3572 \text{ mm Hg}\)
4. \(845 \text{ mm Hg} \times \frac{101.3 \text{ kPa}}{760 \text{ mm Hg}} = 112.6 \text{ kPa}\)

21. c. \(CH_2 = CH_2\) will have the lowest surface tension because there are no groups that could form hydrogen bonds. The other two structures are capable for forming H-bonds.

1. close to 1000 mm Hg.  
2. approximately 170°C

1. The temperature will decrease.  
2. An increase in pressure increases the temperature needed to boil water.
Chapter 14 The Behavior of Gases

Chapter Outline

14.1 Gas Properties
14.2 Gas Laws
14.3 Ideal Gases
14.4 Gas Mixtures and Molecular Speeds
14.5 The Behavior of Gases
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. Most of the volume of a gas is
   1. the gas molecules
   2. rapidly moving particles
   3. empty space
   4. kinetic energy

2. Intermolecular forces in a gas
   1. hold the gas particles together
   2. contribute to gas pressure
   3. slow the particles down
   4. are insignificant

3. Air molecules increase in velocity when
   1. the temperature is raised
   2. when the temperature is lowered
   3. when more molecules are added to the container
   4. when molecules are removed from the container

4. An increase in kinetic energy
   1. increases the volume of a rigid container
   2. increases the pressure in the container
   3. decreases the pressure in the container
   4. decreases the temperature of the system

5. After driving in the desert for several hours, the car tire pressure will
   1. decrease
   2. not change
   3. increase
   4. depends on how fast the car is moving

6. Gas pressure in a rigid vessel will increase when
   1. gas particles are removed
   2. the vessel is cooled
   3. the vessel is heated
   4. the outside pressure is increased

7. One of the following is not part of the kinetic-molecular theory
   1. gas particles move in straight lines
   2. collisions with other particles cause loss of energy
3. Particles do not lose energy when they collide with container walls.
4. Gas particle size is insignificant.

8. An increase in the number of gas particles in a rigid container
   1. Causes a decrease in pressure.
   2. Causes a decrease in temperature.
   3. Increases intermolecular interactions.
   4. Causes an increase in pressure.

9. If the volume of a container of gas is increased, one of the following will not happen
   1. Pressure will decrease.
   2. Number of collisions with a unit area of the container wall will decrease.
   3. Pressure will increase.
   4. Kinetic energy of individual collisions will not be affected.

10. An increase in temperature causes air molecules to
    1. Increase in kinetic energy.
    2. Decrease in kinetic energy.
    3. Increase intermolecular interactions.
    4. Decrease intermolecular interactions.

**True/False:**

11. _____ Basketballs are checked for pressure after a game.
12. _____ The air in a fully inflated bicycle tire is at a lower pressure than the outside air.
13. _____ A given volume of gas has a definite three-dimensional character.
14. _____ Gases are less compressible than liquids.
15. _____ Most of the volume of a gas is made up of the volumes of the gas particles.
16. _____ Removing gas particles from a rigid container lowers the pressure in the system.
17. _____ Doubling the Celsius temperature doubles the pressure of a rigid container.
18. _____ An increase in temperature increases the kinetic energy of gas particles.
19. _____ At STP, the average distance between gas particles is about ten times the diameter of the gas particle.
20. _____ An elastic collision is one in which kinetic energy is lost.

**Fill in the blank:**

21. Gases are unlike other states of _______ in that a gas _______ to fill the shape and _______ of its container.
22. _______ is a measure of how much a given _______ of matter decreases when placed under pressure.
23. If the volume of a container is _________, the gas molecules have less space in which to move around.

**Short Answer:**

24. A balloon is filled to a pressure of 1 atm on the top of Mt. Whitney (14,505 feet) in California during the winter (snow is on the ground and the temperature is about 15°F). The balloon is quickly transported by drone to Badwater in Death Valley (282 feet below sea level), just 76 miles away from Mt. Whitney. The temperature at Badwater is 64°F. Describe the effects of these changes on the balloon.

25. Your friend wants to use an inflatable rubber raft to do a Polar Bear Plunge on New Year’s Day. He inflates the raft to one atm in his garage (74°F). He then drives out to the lake where he and his friends are going to swim (very briefly!) in water at a temperature of 36°F. He hopes the raft will keep him out of the water. What do you think of his idea?
Answer Key

1. c
2. d
3. a
4. b
5. c
6. c
7. b
8. d
9. c
10. a
11. false
12. false
13. false
14. false
15. false
16. true
17. false
18. true
19. true
20. false
21. matter, expands, volume
22. compressibility, volume
23. decreased
24. The balloon will expand in volume. The temperature increase will cause part of the expansion and the increase in pressure when moving to sea level will cause the rest of the expansion.
25. The decrease in temperature will cause a decrease in the volume of the raft. This drop in volume may be enough to sink the passenger when he gets on the raft.
14.2 Gas Laws

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. Boyle’s Law can be expressed mathematically as
   1. \( \frac{P}{V} = k \)
   2. \( P \times V = k \)
   3. \( P \times k = V \)
   4. \( \frac{V}{P} = k \)

2. Boyle’s Law states that, for a gas at constant temperature,
   1. the volume increases as the pressure increases
   2. the volume decreases as the pressure decreases
   3. the volume increases as the pressure decreases
   4. the volume decreases as the pressure decreases

3. Charles’ Law states that, for a gas at constant pressure
   1. the volume increases as the temperature increases
   2. the volume decreases as the temperature increases
   3. the volume increases as the temperature decreases
   4. an increase in volume causes an increase in temperature

4. Charles’ Law can be stated
   1. \( \frac{V}{T} = k \)
   2. \( V \times T = k \)
   3. \( \frac{T}{V} = k \)
   4. \( V \times k = T \)

5. For Gay-Lussac’s Law to be applicable, a _____ container is needed
   1. flexible
   2. large
   3. rigid
   4. small

6. Avogadro’s Law relates
   1. number of moles and temperature
   2. volume and number of moles
   3. number of grams and pressure
   4. number of grams and temperature

7. Gay-Lussac’s Law states
   1. \( \frac{P}{T} = K \)
   2. \( \frac{P}{T} = k \)
3. \( K \times P = T \)
4. \( P \times T = k \)

8. In Avogadro’s Law, if a rigid container is used, the relationship exists between
   1. volume and temperature
   2. pressure and volume
   3. pressure and number of moles
   4. pressure and temperature

9. When using the combined gas law, the only constant is
   1. temperature
   2. amount of gas
   3. pressure
   4. volume

10. The combined gas law contains all the individual gas laws except
   1. Boyle
   2. Charles
   3. Gay-Lussac
   4. Avogadro

**True/False:**

11. _____ In an inverse relationship, as one variable increase, the other also increases.
12. _____ A plot of Charles’ Law gives a straight line.
13. _____ For accurate calculations all temperatures must be in Celsius.
14. _____ As the absolute temperature approaches zero K, the volume also approaches zero.
15. _____ Charles’ Law requires a flexible container.
16. _____ A plot of Boyle’s Law gives a straight line.
17. _____ An increase in temperature produces an increase in kinetic energy.
18. _____ In Avogadro’s Law calculations, grams of material can be used for \( n \).
19. _____ In combined gas law calculations, pressures can only be in atm.
20. _____ The formula for the combined gas law is \( P \times \frac{V}{T} = k \).

**Fill in the Blank:**

21. At constant temperature, \( V_1 = 650 \text{ mL}, P_1 = 465 \text{ torr}, V_2 = 360 \text{ mL}, P_2 =? \)
22. At constant temperature, \( V_1 = 345 \text{ mL}, P_1 = 780 \text{ torr}, P_2 = 970 \text{ torr}, V_2 =? \)
23. At constant pressure, the temperature of 550 mL of gas is increased from 270 K to 420 K. What is the new volume?
24. At constant pressure, the volume of helium increase from 560 mL at 245 K to 970 mL. What is the new temperature?
25. At constant volume, a tank of gas shows a pressure of 2500 atm at 290 K. The tank is then placed in a metal shed where the temperature is 320 K. What is the new pressure?
26. A rigid cylinder of gas shows a pressure of 600 torr at 215 K. It is moved to a new storage site where the pressure is now 750 torr. What is the new temperature?
27. Carry out the following calculations:
28. a. 
29. \( P_1 = 3.5 \text{ atm} \) \( P_2 =? \)
30. \( V_1 = 450 \text{ mL} \) \( V_2 = 790 \text{ mL} \)
31. \( T_1 = 290 \text{ K} \) \( T_2 = 360 \text{ K} \)
32. b.
33. $P_1 = 840 \text{ torr}$ $P_2 = 390 \text{ torr}$
34. $V_1 = 245 \text{ mL}$ $V_2 =$?
35. $T_1 = 940 \text{ K}$ $T_2 = 780 \text{ K}$

1. A 2.4 L balloon holds 3.7 mol He. If 1.6 mol He are added to the balloon, what is the new volume?
2. A balloon containing 2.6 mol hydrogen has a volume of 3.9 L. More hydrogen is added to the balloon, giving it a volume of 17.1 L. How many moles of hydrogen were added?

Answer Key

1. b
2. c
3. a
4. a
5. c
6. b
7. a
8. c
9. b
10. d
11. false
12. true
13. true
14. false
15. true
16. false
17. true
18. false
19. false
20. true

21.

$$P_1V_1 = P_2V_2$$

$$P_2 = \frac{P_1V_1}{V_2}$$

$$= \frac{465 \text{ torr} \times 650 \text{ mL}}{360 \text{ mL}}$$

$$= 839.6 \text{ torr}$$

22.

$$V_2 = \frac{P_1V_1}{P_2}$$

$$= \frac{780 \text{ torr} \times 345 \text{ mL}}{970 \text{ mL}}$$

$$= 277.4 \text{ mL}$$
23. \[
\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ or } V_2 = \frac{V_1 T_2}{T_1} \\
= \frac{550 \text{ mL} \times 420 \text{ K}}{270 \text{ K}} \\
= 855.6 \text{ mL}
\]

24. \[
\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ or } T_2 = \frac{T_1 V_2}{V_1} \\
= \frac{245 \text{ K} \times 970 \text{ mL}}{560 \text{ mL}} \\
= 424.4 \text{ mL}
\]

25. \[
\frac{P_1}{T_1} = \frac{P_2}{T_2} \text{ or } P_2 = \frac{P_1 T_2}{T_1} \\
= \frac{2500 \text{ atm} \times 320 \text{ K}}{290 \text{ K}} \\
= 2758 \text{ atm}
\]

26. \[
\frac{P_1}{T_1} = \frac{P_2}{T_2} \text{ or } T_2 = \frac{P_2 T_1}{P_1} \\
= \frac{750 \text{ torr} \times 215 \text{ K}}{600 \text{ torr}} \\
= 268.8 \text{ K}
\]

1. \[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \text{ or } P_2 = \frac{P_1 V_1 T_2}{T_1 V_2} = \frac{3.5 \text{ atm} \times 450 \text{ mL} \times 360 \text{ K}}{290 \text{ K} \times 790 \text{ mL}} = 2.47 \text{ atm}
\]

2. \[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \text{ or } V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{840 \text{ torr} \times 245 \text{ mL} \times 780 \text{ K}}{940 \text{ K} \times 390 \text{ torr}} = 437.9 \text{ mL}
\]

27. a. 

28. \[n_1 = 3.7 \text{ mol } n_2 = 3.7 + 1.6 = 5.3 \text{ mol}\]

29. \[
\frac{V_1}{n_1} = \frac{V_2}{n_2} \text{ or } V_2 = \frac{V_1 n_2}{n_1} = \frac{2.4 \text{ L} \times 5.3 \text{ mol}}{3.7 \text{ mol}} = 3.44 \text{ L}
\]

30. b. find \(n\)

31. 2

32. added moles = \(n_2 - n_1\)

33. \[
\frac{V_1}{n_1} = \frac{V_2}{n_2} \text{ or } n_2 = \frac{V_1 n_1}{V_2} = \frac{17.1 \text{ L} \times 2.6 \text{ mol}}{3.9 \text{ L}} = 11.4 \text{ mol}
\]

34. added moles = 11.4 - 2.6 = 8.8 moles
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. The ideal gas law is a combination of
   1. Boyle’s law and Charles’ law
   2. Avogadro’s law and Gay-Lussac’s law
   3. combined gas law and Avogadro’s law
   4. combined gas law and law of conservation of mass

2. The ideal gas law is stated
   1. \[ PT = nRV \]
   2. \[ PV = nRT \]
   3. \[ PV = \frac{RT}{n} \]
   4. \[ n = \frac{RT}{PV} \]

3. The units of R require all of the following except
   1. volume be in liters
   2. temperature be in Kelvin
   3. pressure be in mm Hg, kPa, or atm
   4. pressure be in mm Hg, atm, or torr

4. The unit for R is \( J/K \cdot mol \) when the pressure is in
   1. mm Hg
   2. kPa
   3. atm
   4. psi

5. The units for \( n \) are
   1. grams
   2. mass
   3. moles
   4. molar mass

6. Combustion reactions involve all of the following except
   1. oxygen as a reactant
   2. carbon dioxide as a product
   3. water as a product
   4. nitrogen as a reactant

7. The ideal gas law allows stoichiometry calculations
   1. at any temperature and pressure
   2. using unbalanced equations
3. at standard pressure only
4. at standard temperature only

8. One of the following statements about ideal gases is not true
   1. \( \frac{PV}{RT} \) ratio very close to one at all temperatures
   2. particles closely packed
   3. negligible intermolecular interactions
   4. size of particles insignificant

9. At higher pressures, \( \frac{PV}{RT} \) becomes less than one because
   1. attractive forces become more significant
   2. the space occupied by the particles becomes more significant
   3. the kinetic energy of gas particles becomes less
   4. the temperature increases

10. At the same temperature and pressure, which of the following gases is least likely to demonstrate ideal gas behavior?
    1. O\(_2\)
    2. Ne
    3. H\(_2\)O
    4. H\(_2\)

True/False:

11. _____ For an ideal gas, the plot of \( \frac{PV}{RT} \) vs. pressure is a horizontal line.
12. _____ The ideal gas law can only be used at STP.
13. _____ Temperature units for the ideal gas law can be either Celsius or Fahrenheit.
14. _____ \( kPa \times L = J \)
15. _____ Only temperature and pressure can be calculated using the ideal gas law.
16. _____ Avogadro’s Law shows that pressure is directly proportional to \( n \).
17. _____ The ideal gas law shows that an increase in volume produces a decrease in pressure at constant \( n \) and \( T \).
18. _____ The numerical value of \( R \) depends on the units for temperature.
19. _____ Evaluation of \( R \) requires that the gas be ideal.
20. _____ Gas density at non-standard conditions cannot be determined using the ideal gas law.

Fill in the Blank:

21. The combined gas law shows that the pressure of a gas is ________ proportional to volume and directly proportional to _________.
22. The ________ \( R \) in the equation is called the ideal gas _________.
23. An ideal gas is one that follows the ____ laws at all conditions of ______ and ______.
24. HF can exist as a gas above 293 K and one atm pressure. Would you expect HF to behave as an ideal gas? Explain your answer.
25. What volume of gas is occupied by 4.7 moles of nitrogen at 675 K and a pressure of 3.9 kPa?
26. How many moles of Rn are present in a sample that occupies a volume of 125 L at a pressure of 0.75 kPa and a temperature of 305 K?
27. Ammonium nitrate is heated to form N\(_2\)O and water according to the following equation:
   \( NH_4NO_3 \rightarrow N_2O + 2H_2O \)
28. What mass of ammonium nitrate needs to be heated to form 125 mL of N\(_2\)O?
29. At 125 kPa and 300 K?
30. s
31. O at 125 kPa and 300 K?
32. A gas sample occupies a volume of 150 mL at a temperature of 295 K and a pressure of 98.2 kPa. How many moles of gas are there?
33. A gas occupies a space of 2.9 L at a pressure of 3.6 atm and a temperature of 315 K. It has a mass of 24.2 g. What is the molar mass?

34. Calculate the density of neon gas at 350 K and a pressure of 845 mm Hg.

---

**Answer Key**

1. c  
2. b  
3. d  
4. b  
5. c  
6. d  
7. a  
8. b  
9. a  
10. c  
11. true  
12. false  
13. false  
14. true  
15. false  
16. true  
17. true  
18. false  
19. true  
20. false  
21. inversely, temperature  
22. variable, constant  
23. gas, temperature, pressure  
24. No, HF would not behave as an ideal gas. The electronegativity of F is strong enough to create significant intermolecular interactions among HF molecules.

25. \( PV = nRT \) and \( V = \frac{nRT}{P} \)

26. \( V = \frac{4.7 \text{ mol} \times 8.314 \text{ J/K mol} \times 675 \text{ K}}{3.9 \text{ kPa}} = 6763 \text{ L} \)

27. \( PV = nRT \) and \( n = \frac{PV}{RT} \)

28. \( n = \frac{0.75 \text{ kPa} \times 125 \text{ L}}{8.314 \text{ J/K mol} \times 305 \text{ K}} = 0.037 \text{ mol} \)

29. \( n = \frac{PV}{RT} \)

30. \( n = \frac{125 \text{ kPa} \times 125 \text{ L}}{8.314 \text{ J/K mol} \times 300 \text{ K}} = 6.26 \text{ mol} \text{ N}_2\text{O} \)

31. 1:1 ratio of moles ammonium nitrate to moles N

32. 2

33. O molar mass of ammonium nitrate is

34. 

\[ N_2 \times 14.0 = 28 \]
\[ H_4 \times 1.01 = 4.04 \]
\[ O_3 \times 16 = 48 \]
\[ 80 \text{ g/mol} \]

35. \( 80 \text{ g/mol} \times 6.26 \text{ mol} = 501 \text{ g ammonium nitrate} \)
36.

\[
n = \frac{PV}{RT} = \frac{98.2 \text{ kPa} \times 150 \text{ L}}{8.314 \text{ J/K mol} \times 295 \text{ K}} = 6.01 \text{ mol}
\]

37.

\[
n = \frac{PV}{RT} = \frac{3.6 \text{ atm} \times 2.9 \text{ L}}{0.08206 \text{ L atm/K mol} \times 315 \text{ K}} = 0.40 \text{ mol}
\]

38. \(\frac{24.2 \text{ g}}{0.40 \text{ mol}} = 60.5 \text{ g/mol}\)

39. molar mass of neon = 20.2 g/mol

40. \(V = \frac{nRT}{P} = \frac{1.0 \text{ mol} \times 62.36 \text{ L mm Hg/K mol} \times 350 \text{ K}}{845 \text{ mm Hg}} = 28.8 \text{ L}\)

41. density = \(\frac{\rho}{L} = \frac{20.2}{28.8} = 0.70 \text{ g/L}\)
Lesson Quiz

Multiple Choice:

1. The partial pressure of neon is indicated by
   1. \( P_{\text{ne}} \)
   2. \( P_{\text{Ne}} \)
   3. \( p_{\text{Ne}} \)
   4. partial \( P_{\text{Ne}} \)

2. The symbol for mole fraction is
   1. \( m_f \)
   2. \( X \)
   3. \( F_{\text{mol}} \)
   4. \( x \)

3. The partial pressure of a gas equals
   1. mole fraction \( \times \) total pressure
   2. mole fraction \( \div \) total pressure
   3. total pressure \( \div \) mole fraction
   4. \( (\text{total pressure})^2 \times \text{mole fraction} \)

4. Water displacement involves
   1. collecting a gas by bubbling it through water
   2. bubbling water through the gas
   3. adding water to the gas after it is collected
   4. removing water from the gas while it is being produced

5. Gas collected by water displacement will
   1. be contaminated by air
   2. dissolve in the water
   3. contain water vapor
   4. be pure

6. The pressure of a gas containing water vapor is an application of
   1. Charles’ Law
   2. Dalton’s Law
   3. Boyle’s Law
   4. Gay-Lussac’s Law

7. Graham’s Law associates gas diffusion with
   1. pressure
   2. volume
   3. actual mass
   4. molar mass
8. If one mole of $N_2$ at 760 torr is added to a container holding one mole of $H_2$ at 760 torr
   1. the pressure in the container is 1520 torr
   2. the pressure in the container is 380 torr
   3. the pressure in the container is 760 torr
   4. the pressure in the container is 1140 torr

9. The mole fraction of nitrogen in air is calculated using the following formula
   1. $\text{mol } N_2 \times \text{total mol}$
   2. $\text{mol } N_2 \div \text{total mol}$
   3. $\text{total mol} \div \text{mol } N_2$
   4. $(\text{mol } N_2 \times \text{total mol}) \times 100$

10. The partial pressure of a gas ($P_x$) in a mixture can be determined with
    1. $P_x = \text{mole fraction}_x \div \text{total moles}$
    2. $P_x = \text{total moles} \div \text{mole fraction}_x$
    3. $P_x = \text{mole fraction}_x \times \text{total pressure}$
    4. $P_x = \text{mole fraction}_x \div \text{total pressure}$

True/False:

11. _____ The total pressure in a mixture of gases is the sum of the partial pressures of the gases.
12. _____ A gas collected over water contains water vapor due to boiling at a high temperature.
13. _____ The vapor pressure of water is not affected by temperature.
14. _____ Carbon dioxide in the atmosphere exerts more pressure in air than does oxygen.
15. _____ Solids diffuse less readily than liquids.
16. _____ At the same temperature and volume, a container holding 1200 gas particles exhibits more pressure
    than a container holding 745 particles.
17. _____ Graham’s Law applies to both diffusion and effusion.
18. _____ Dalton found that the pressure of a gas in a mixture is affected by other gases in the system.
19. _____ The mole fraction of a gas in a mixture depends upon its structure.
20. _____ Gases with high molar masses diffuse more slowly than gases with low molar masses.

Fill in the Blank:

21. Define the following terms:
    1. Dalton’s Law
    2. Graham’s Law
    3. partial pressure
    4. diffusion
22. Which of the following gases will diffuse the slowest? the fastest?
23. $N_2, He, HC \equiv CH, Ar, Cl_2$
24. Determine the partial pressure of oxygen in a container where it is mixed with hydrogen (partial pressure 85
    mm Hg) and helium (partial pressure 198 mm Hg). The pressure in the tank is 412 mm Hg.
25. Space capsules operate with an oxygen content of about 34%. Assuming a total pressure of 780 mm Hg in the
    space capsule, what is the partial pressure of the oxygen?
26. Under a given set of circumstances, He effuses at a rate of 425 m/s. What is the effusion rate for Ar under
    those circumstances?
27. Calculate the mole fraction of each gas in a mixture of 43.7 g Xe and 36.9 g Kr.
28. If the total pressure in the system is 267 kPa, what is the partial pressure of the Xe and Kr in problem 26?
29. An experiment generates 25.0 L of O$_2$ collected over water. The atmospheric pressure is 740 mm Hg and the
    temperature is 25°C. What is the volume of the dry oxygen at STP?
1. The total pressure of a mixture of gases is equal to the sum of all of the partial pressures of the component gases.
2. The rate of effusion or diffusion of a gas is inversely proportional to the square root of the molar mass of the gas.
3. The contribution that one gas makes to the total pressure when the gas is part of a mixture.
4. The tendency of molecules to move from an area of high concentration to an area of low concentration until the concentration is uniform.
Chapter Test

Multiple Choice:

35. Car tire pressure is adjusted in a warm garage and the car is then driven for several hours in subzero weather. The tire pressure will
   1. decrease
   2. not change
   3. increase
   4. depends on how fast the car is moving

2. The gas law that states \( PV = k \) is
   1. Boyle’s Law
   2. Charles’ Law
   3. Dalton’s Law
   4. Gay-Lussac’s Law

3. \( PV = nRT \) is a statement of
   1. Dalton’s Law
   2. law of partial pressures
   3. ideal gas law
   4. Avogadro’s Law

4. Mole fraction \( \times \) total pressure is a way to calculate
   1. molar mass
   2. partial pressures
   3. universal gas constant
   4. vapor pressure

5. An ideal gas has the following property
   1. particles are close together
   2. molar mass affects pressure
   3. intermolecular interactions negligible
   4. particle size important for pressure

6. Which of the following gases is least likely to show ideal behavior?
   1. HF
   2. Ne
   3. \( \text{Cl}_2 \)
   4. \( \text{O}_2 \)

7. When gas pressure is measured in kPa, the units for \( R \) are
   1. \( L\cdot mm\ Hg/K\cdot mol \)
   2. \( J/K\cdot mol \)
   3. \( L\cdot atm/K\cdot mol \)
4. \( kPa/K \cdot mol \)

8. The law that relates volume and number of moles of a gas is
   1. Dalton’s Law
   2. Avogadro’s Law
   3. Charles’ Law
   4. Boyle’s Law

9. A decrease in temperature causes air molecules to
   1. increase in kinetic energy
   2. increase intermolecular interactions
   3. decrease intermolecular interactions
   4. decrease in kinetic energy

10. The combination of the combined gas law and Avogadro’s law is
   1. partial pressures law
   2. law of conservation of mass
   3. ideal gas law
   4. molecular mass law

True/False:

11. _____ The plot of \( PV \over RT \) for an ideal gas is a curved line.
12. _____ At the same temperature and volume, a container holding 450 gas particles exhibits more pressure than a container holding 345 particles.
13. _____ Graham’s Law is derived from an equation for kinetic energy.
14. _____ The ideal gas law shows that an increase in volume produces an increase in pressure at constant \( n \) and \( T \).
15. _____ A plot of Boyle’s Law gives a straight line.
16. _____ Doubling the Kelvin temperature doubles the pressure in a rigid container.
17. _____ A decrease in temperature increases the kinetic energy of gas particles.
18. _____ An elastic collision is one in which kinetic energy is gained.
19. _____ The amount of gas is assumed to be constant when doing combined gas law calculations.
20. _____ Temperature units for gas law calculations are in Kelvin.

Fill in the Blank:

21. Define the following terms:
   1. diffusion
   2. partial pressure
   3. Avogadro’s Law
   4. Charles’ Law

22. Make the following calculations:
   1. \( P_1 = 3.6 \text{ atm} \ V_1 = 246 \text{ L} \ T_1 = 375 \text{ K} \ \ \ \ P_2 = ? \ \ V_2 = 816 \text{ L} \ T_2 = 523 \text{ K} \)
   2. \( P_1 = 840 \text{ mm Hg} \ V_1 = 241 \text{ L} \ T_1 = 740 \text{ K} \ P_2 = 980 \text{ mm Hg} \ V_2 = ? \ T_2 = 460 \text{ K} \)

23. Calculate the mole fraction for each gas in a mixture of 46.7 g Kr and 6.9 g H\(_2\).

24. A balloon containing 4.5 moles of He has a volume of 23.1 L. More helium is added, bringing the balloon to a volume of 35.2 L. How many moles He were added?

25. What volume of gas is occupied by 6.4 moles of oxygen at 298 K and a pressure of 840 mm Hg?

26. A gas occupies a volume of 4.9 L at a pressure of 502 kPa and a temperature of 356 K. It has a mass of 36.5 g. What is the molar mass?
27. Determine the partial pressure of He in a tank containing argon (partial pressure 340 mm Hg) and neon (partial pressure 195 mm Hg). The total pressure in the tank is 850 mm Hg.

28. An experiment produces 33 L of H₂ collected over water. The atmospheric pressure is 755 mm Hg and the temperature is 20°C. What is the volume of the dry hydrogen at STP?

---

**Answer Key**

1. a
2. a
3. c
4. b
5. c
6. a
7. b
8. b
9. d
10. c
11. false
12. true
13. true
14. false
15. false
16. true
17. false
18. false
19. true
20. true

1. The tendency of molecules to move from an area of high concentration to an area of low concentration until the concentration is uniform.
2. The contribution that one gas makes to the total pressure when the gas is part of a mixture.
3. The volume of a gas is directly proportional to the number of moles of gas when the temperature and pressure are held constant.
4. The volume of a given mass of gas varies directly with the absolute temperature of the gas when the pressure is kept constant.

1. \( \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \) or \( P_2 = \frac{P_1 V_1 T_2}{T_1 V_2} = \frac{3.6 \text{ atm} \times 246 \text{ mL} \times 523 \text{ K}}{375 \text{ K} \times 816 \text{ mL}} = 1.51 \text{ atm} \)
2. \( \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \) or \( V_2 = \frac{P_1 V_1 T_2}{T_1 P_2} = \frac{840 \text{ mm Hg} \times 241 \text{ L} \times 460 \text{ K}}{740 \text{ K} \times 960 \text{ mm Hg}} = 128.4 \text{ L} \)

21. 46.7 g ÷ 83.8 g/mol = 0.56 mol Kr  6.9 g ÷ 2.02 g/mol = 3.42 mol H₂

22. total moles = 0.56 + 3.42 = 3.98 mol

23. \( X_{\text{krpton}} = 0.56 ÷ 3.98 = 0.14 \times \)

hydrogen = 3.42 ÷ 3.98 = 0.86

24. \( \frac{V_1}{n_1} = \frac{V_2}{n_2} \) or \( n_2 = \frac{V_2 n_1}{V_1} = \frac{35.2 \text{ L} \times 4.5 \text{ mol}}{23.1 \text{ L}} = 6.86 \text{ mol} \)

25. 6.86 mol – 4.5 mol = 2.36 mole added

26. 

27. \( PV = nRT \) or \( V = \frac{nRT}{P} \)  \( V = \frac{6.4 \text{ mol} \times 62.36 \text{ L mm Hg} / \text{K} - \text{mol} \times 298 \text{ K}}{840 \text{ mm Hg}} = 141.6 \text{ L} \)

28. \( PV = nRT \) or \( n = \frac{PV}{RT} \)  \( n = \frac{4.9 \text{ L} \times 502 \text{ kPa}}{314 \text{ J/K/mol} \times 356 \text{ K}} = 0.83 \text{ mol} \)

29. 36.5 g ÷ 0.83 mol = 44 g/mol

30. \( P_{\text{total}} = P_{\text{He}} + P_{\text{Ar}} + P_{\text{Ne}} \)
31. \( P_{He} = P_{\text{total}} - (P_{Ar} + P_{Ne}) = 850 \text{ mm Hg} - (340 + 195) = 850 - 535 = 315 \text{ mm Hg} \)

32. \( 755 \text{ mm Hg} - 17.5 \text{ mm Hg} = 737.5 \text{ mm Hg} \) dry pressure

33. \( V_2 = \frac{P_1 \times V_1 \times T_2}{P_2 \times T_1} = \frac{737.5 \text{ mm Hg} \times 33 \text{ L} \times 273 \text{ K}}{760 \text{ mm Hg} \times 293 \text{ K}} = 29.8 \text{ L} \)
Chapter 15

Water Assessments

Chapter Outline

15.1 Properties of Water
15.2 Aqueous Solutions
15.3 Colloids and Suspensions
15.4 Water
Lesson Quiz

1. Water exhibits a bent molecule configuration because of
   1. attractions between O and H
   2. sp³ hybridization
   3. two sets of lone pair O electrons
   4. repulsions between the two δ⁺H atoms

2. The H-O-H bond angle is
   1. 109.5°
   2. 105°
   3. 109.47°
   4. 107°

3. Each oxygen atom has a ________ geometry around it.
   1. square planar
   2. trigonal
   3. orthorhombic
   4. tetrahedral

4. Ice floats because
   1. there is more space between molecules
   2. repulsion by partial negative charge on O
   3. there is less space between molecules
   4. repulsion by partial positive charge on H

5. The bent shape of the water molecule
   1. allows molecules to come closer together
   2. makes the entire molecule polar
   3. permits hydrogen bond formation
   4. helps in dissolving materials

6. The O-H bond is polar covalent because of
   1. the bent shape of the molecule
   2. the electronegativity of the H atom
   3. the lone pair electrons on O
   4. the electronegativity of the O atom

7. Hydrogen bonds are stronger than
   1. covalent bonds
   2. ionic bonds
3. ion-dipole bonds  
4. dipole-dipole bonds  

8. The high boiling point of water is due to  
   1. the bent shape of the water molecule  
   2. hydrogen bonding  
   3. dipole-dipole interactions  
   4. kinetic energy of the water molecule  

9. The most accurate meniscus reading is  
   1. at eye level  
   2. below the level of the meniscus  
   3. above the level of the meniscus  
   4. taking an average of readings at different levels  

10. The low vapor pressure of water is due to  
    1. high kinetic energy of the water molecules  
    2. hydrogen bonding  
    3. high molecular mass of the water molecule  
    4. dipole-dipole interactions  

True/False:  
11. _____ When water is in a thin column, the center of the surface is lower than the edges.  
12. _____ The low vapor pressure of water decreases the rate of evaporation from bodies of water.  
13. _____ Water reaches its highest density at 2°C.  
14. _____ There are many solids that are less dense than their liquid form.  
15. _____ As water approaches 0°C, the molecules begin to organize into a octahedral structure.  
16. _____ The bent shape of the water molecule makes the entire structure polar.  
17. _____ Each H atom can form a hydrogen bond with one O atom.  
18. _____ In a water molecule, the H atoms function as dipoles.  
19. _____ Very little energy is required to break the H-bonds in ice.  
20. _____ As water cools, its molecular motion slows.  

Fill in the Blank:  
21. Oceans, rivers, and lakes cover about ____ of the Earth’s surface.  
22. The highly _____ O-H bonds leave very little _____ density around the ____ atoms.  
23. The geometry around each oxygen atom consists of two _______ bonds and two _______ bonds.  
24. The _______ of any liquid increases as its temperature _______.  
25. _____ is one of only very few solids that is _____ dense than its liquid form.  
26. Look at the water cycle diagram at the beginning of the lesson. List the ways that liquid water can gain access to the atmosphere.

Answer Key  
1. c  
2. b  
3. d  
4. a
5. b  
6. d  
7. d  
8. b  
9. a  
10. b  
11. true  
12. true  
13. false  
14. false  
15. false  
16. true  
17. true  
18. true  
19. false  
20. true  
21. 75%  
22. polar, electron, hydrogen  
23. covalent, hydrogen  
24. density, decreases  
25. ice, less  
26. evaporation, sublimation, evapotranspiration
15.2 Aqueous Solutions

Lesson Quiz

Name___________________ Class______________ Date________

Multiple Choice:

1. When NaHCO₃ dissociates in water, the products are
   1. Na⁺ + H⁺ + CO₃²⁻
   2. Na⁺ + H⁺ + CO₃⁻
   3. Na⁺ + HCO₃⁻
   4. Na⁺ + HCO₃²⁻

2. Oil and water form two layers when mixed together. Oil is a ______ material.
   1. polar
   2. nonpolar
   3. ionic
   4. dipole

3. An electrolyte shows all the following properties except
   1. can conduct electricity when dissolved in water or melted
   2. is ionic
   3. does not dissociate into ions when dissolved in water or melted
   4. forms H bonds with water molecules

4. Fe₂(SO₄)₃ is slightly soluble in water and forms _____ ions
   1. 3
   2. 5
   3. 12
   4. 2

5. The double arrow in an ionization equation indicate
   1. ionization of a weak electrolyte
   2. ionization of a strong electrolyte
   3. an incomplete reaction
   4. reaction of a nonelectrolyte

6. Water will dissolve all of the following except
   1. table salt
   2. sugar
   3. vegetable oil
   4. baking soda

7. The geological process of _______ is a direct result of water’s dissolving power
   1. earthquake
   2. typhoon
3. tornado
4. erosion

8. When a crystal of sodium chloride is placed into water,
   1. water molecules are surrounded by sodium ions
   2. water molecules collide with the crystal lattice
   3. the ions react with water molecules
   4. water molecules are surrounded by chloride ions

9. Aqueous solutions are stabilized by
   1. hydration
   2. hydrolysis
   3. precipitation
   4. hyperactivity

10. The attractive forces between nonpolar particles are
    1. ionic forces
    2. covalent forces
    3. dispersion forces
    4. ion-dipole forces

True/False:

11. _____ All ionic compounds are electrolytes.
12. _____ Solutes in a true solution can be filtered out.
13. _____ Ethanol is a polar liquid.
14. _____ When KBr dissolves, the bromine ion is attracted to the O atom in water
15. _____ Sugar dissolves in water to produce ions.
16. _____ HMnO₄ dissociates in water to produce H⁺ and MnO₄⁻ ions.
17. _____ HCl gas is an electrolyte.
18. _____ Solvation involves formation of solutions by nonelectrolytes.
19. _____ Water molecules surround solute particles in a random manner.
20. _____ Nonpolar particles are held together by weak dispersion forces.

Fill in the Blank:

21. An _______ solution is water that contains one or more _______ substances.
22. Liquids that do not _______ in one another are called _______.
23. Write dissociation equations for the following materials. Use NR if the material does not dissociate in water.
   1. Na₂CO₃
   2. CH₄
   3. LiC₂H₃O₂
   4. CH₃CH₂OH
   5. KOH
24. A solution is a _______ mixture consisting of a _______ dissolved into a _______.

Answer Key

1. c
2. b  
3. c  
4. b  
5. a  
6. c  
7. d  
8. b  
9. a  
10. c  
11. true  
12. false  
13. true  
14. false  
15. false  
16. true  
17. false  
18. false  
19. false  
20. true  
21. aqueous, dissolved  
22. dissolve, immiscible  
   1. $Na_2CO_3 \rightarrow 2Na^+ (aq) + CO_3^{2-} (aq)$  
   2. $CH_4 \rightarrow NR$  
   3. $LiC_2H_2O_2 \rightarrow Li^+ (aq) + C_2H_3O_2^- (aq)$  
   4. $CH_3CH_2OH \rightarrow CH_3CH_2OH(aq)$  
   5. $KOH \rightarrow K^+(aq) + OH^- (aq)$  
23. homogeneous, solute, solvent
Lesson Quiz

Name___________________ Class______________ Date________

Multiple Choice:

1. A mixture in which some of the material settles out is called a
   1. precipitate
   2. suspension
   3. colloid
   4. solution

2. One of the following is not a property of a solution
   1. homogeneous mixture
   2. particle size less than 1 nm
   3. can filter material out of solution
   4. does not scatter light

3. Large particles are characteristic of a
   1. suspension
   2. colloid
   3. solution
   4. solid

4. All of the following are colloids except
   1. smoke
   2. butter
   3. milk
   4. cola drink

5. The scattering of light by a colloid is called the _____ effect
   1. Timpani
   2. Thomas
   3. Tyndall
   4. Tyson

6. Colloids exhibit _____ motion, seen under a microscope
   1. Bragton
   2. Brownian
   3. Broenian
   4. Bigstein

7. An example of a solid emulsion colloid is
   1. mayonnaise
   2. butter
3. marshmallow
4. blood

8. A sol and gel colloid consists of
   1. solid dispersed in liquid.
   2. liquid dispersed in liquid
   3. liquid dispersed in solid
   4. gas dispersed in solid

9. Egg yolk is added as ______ to a mixture of oil and vinegar
   1. dispersing agent
   2. emulsifying agent
   3. emulsion agent
   4. solubilizing agent

10. A liquid dispersed in a solid is called a ______ colloid
    1. solid aerosol
    2. liquid aerosol
    3. solid emulsion
    4. liquid emulsion

**True/False:**

11. _____ Colloids move in rapid and random motions
12. _____ Both ends of a soap molecule are polar.
13. _____ Light passed through a solution is not deflected.
14. _____ A colloid is a homogeneous mixture.
15. _____ Suspensions do not stay dispersed unless they are actively mixed.
16. _____ Dispersed colloid particles are about the same size as those in solutions.
17. _____ Suspensions do not scatter light.
18. _____ Mud is an example of a solid emulsion colloid.
19. _____ The rapid movements of colloidal particles can only be seen under a microscope.
20. _____ Some suspensions will be opaque.

**Fill in the Blank:**

21. Define the following terms:
    1. colloid
    2. emulsion
    3. suspension
    4. Tyndall effect

22. Colloidal particles are spread evenly throughout the ______ medium, which can be a _____, ______, or ______.

23. A _____ emulsion requires an _____ agent to be present.

24. Fill in the blanks:

**Table 15.1:**

<table>
<thead>
<tr>
<th>Solutions</th>
<th>Colloids</th>
<th>Suspensions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Homogeneous</td>
<td>Particle size: 1-1000 nm, dispersed; large molecules or aggregates</td>
<td></td>
</tr>
</tbody>
</table>
Table 15.1: (continued)

<table>
<thead>
<tr>
<th>Solutions</th>
<th>Colloids</th>
<th>Suspensions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Particles settle out</td>
<td>Cannot be separated by filtration</td>
<td>May either scatter light or be opaque</td>
</tr>
<tr>
<td>Do not scatter light</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Answer Key**

1. b
2. c
3. a
4. d
5. c
6. b
7. b
8. a
9. b
10. c
11. true
12. false
13. true
14. false
15. true
16. false
17. false
18. false
19. true
20. true

1. a heterogeneous mixture in which the dispersed particles are intermediate in size between those of a solution and a suspension.
2. a colloidal dispersion of a liquid in either a liquid or a solid.
3. a heterogeneous mixture in which some of the particles settle out of the mixture upon standing.
4. the scattering of visible light by colloidal particles.

21. dispersion, solid, liquid, gas.
22. stable, emulsifying

Table 15.2:

<table>
<thead>
<tr>
<th>Solutions</th>
<th>Colloids</th>
<th>Suspensions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Homogeneous</td>
<td>Heterogeneous</td>
<td>Heterogeneous</td>
</tr>
<tr>
<td>Particle size: 0.01-1 nm; atoms, ions, or molecules</td>
<td>Particle size: 1-1000 nm, dispersed; large molecules or aggregates</td>
<td>Particle size: over 1000 nm, suspended; large particles or aggregates</td>
</tr>
<tr>
<td>Do not separate on standing</td>
<td>Do not separate on standing</td>
<td>Particles settle out</td>
</tr>
<tr>
<td>Cannot be separated by filtration</td>
<td>Cannot be separated by filtration</td>
<td>Can be separated by filtration</td>
</tr>
<tr>
<td>Solutions</td>
<td>Colloids</td>
<td>Suspensions</td>
</tr>
<tr>
<td>-------------------</td>
<td>--------------------------------</td>
<td>------------------------------------------</td>
</tr>
<tr>
<td>Do not scatter light</td>
<td>Scatter light (Tyndall effect)</td>
<td>May either scatter light or be opaque</td>
</tr>
</tbody>
</table>
Chapter Test

Name___________________ Class________________ Date________

Multiple Choice:

1. Acetone and water mix together because acetone is a ______ molecule.
   1. polar
   2. nonpolar
   3. ionic
   4. dipole

2. Water exhibits a ______ molecular configuration.
   1. linear
   2. bent
   3. tetrahedral
   4. trigonal

3. The shape of the water molecule
   1. facilitates solution formation
   2. allows colloids to form
   3. makes the entire molecule polar
   4. helps in hydrogen bond formation

4. Motor oil forms a colloid with water only if aided by
   1. a solubilizer
   2. increased temperature
   3. an emulsifying agent
   4. egg yolk

5. The Tyndall effect is
   1. visualization of a suspension
   2. light scattering by a colloid
   3. light scattering by a solution
   4. rapid movement of colloid particles

6. KClO₄ dissociates to form
   1. \( K^+ + ClO_4^- \)
   2. \( K^+ + Cl^- + 2O_2 \)
   3. \( KCl + 2O_2 \)
   4. \( KClO_2 + O_2 \)

7. Hydrogen bonding affects all of the following properties of water except
   1. solubility of polar materials
   2. high melting point
3. low boiling point
4. lower density of solid water

8. The meniscus for water in a thin glass column is due to
   1. low polarity
   2. high surface tension
   3. repulsion of water by the glass
   4. molecular mass of water

9. LiCl dissolves in water because
   1. water molecules are surrounded by lithium ions
   2. water molecules collide with the crystal lattice
   3. the ions react with water molecules
   4. water molecules are surrounded by chloride ions

10. Shaving cream is an example of a _______ colloid.
    1. sol and gel
    2. solid aerosol
    3. liquid aerosol
    4. foam

True/False:

11. _____ The double arrow in an ionization equation indicates ionization of a weak electrolyte.
12. _____ Dipole-dipole interactions are responsible for the low vapor pressure of water.
13. _____ Each H atom can form H-bonds with two O atoms.
14. _____ Hydrogen bonds are stronger than ionic bonds.
15. _____ The most accurate meniscus readings are made at eye level.
16. _____ Ice is less dense than water.
17. _____ Water will not dissolve CH₃OH.
18. _____ Non-polar molecules interact by way of dispersion forces.
19. _____ Electrolytes dissolve by forming H-bonds with water.
20. _____ Paint is a sol and gel colloid.

Fill in the Blank:

21. Define the following terms:
    1. hydration
    2. miscible
    3. solute
    4. electrolyte
    5. solvation

22. A _______ emulsion requires an _______ agent to be present.
23. The geometry around each oxygen atom consists of two _______ bonds and two _______ bonds.
24. Write dissociation reactions for the following:
    1. NaBr (strong electrolyte)
    2. HCN (weak electrolyte)
    3. HCl (strong electrolyte)
    4. HF (weak electrolyte)
**Answer Key**

1. a  
2. b  
3. c  
4. c  
5. b  
6. a  
7. c  
8. b  
9. b  
10. d  
11. true  
12. false  
13. false  
14. false  
15. true  
16. true  
17. false  
18. true  
19. false  
20. true

1. the process of solute particles being surrounded by water molecules arranged in a specific manner.  
2. liquids that dissolve in one another in all proportions.  
3. the substance that is being dissolved  
4. a compound that conducts an electric current when it is dissolved in water or melted.  
5. individual ions are surrounded by solvent particles.

21. colloidal, emulsifying  
22. covalent, hydrogen

1. \( \text{NaBr}(aq) \rightarrow \text{Na}^+(aq) + \text{Br}^-(aq) \)  
2. \( \text{HCN}(aq) \leftrightharpoons \text{H}^+(aq) + \text{CN}^-(aq) \)  
3. \( \text{HCl}(aq) \rightarrow \text{H}^+ + \text{Cl}^- \)  
4. \( \text{HF}(aq) \leftrightharpoons \text{H}^+(aq) + \text{F}^-(aq) \)
Lesson Quiz

Multiple Choice:

1. Some liquid medications can be administered using an oxygen inhaler. This is an example of a ________-solution.
   1. gas in liquid
   2. gas in gas
   3. liquid in gas
   4. liquid in liquid

2. An example of a solid in liquid solution is
   1. water in ammonia
   2. cocoa in water
   3. gravel in concrete
   4. ice in water

3. The form of salt that will dissolve least rapidly is
   1. salt block
   2. salt crystals in rock salt
   3. salt from the salt shaker
   4. salt for melting ice

4. Sugar will dissolve most quickly if the solution is
   1. allowed to sit
   2. stirred rapidly
   3. stirred slowly
   4. chilled before stirring

5. NaCl is dissolved in water until no more will go into solution. At this point the solution is
   1. saturated
   2. unsaturated
   3. supersaturated
   4. partially saturated

6. Materials in aqueous solution
   1. surround water molecules
   2. are hydrated by water molecules
   3. form hydrogen bonds with water molecules
   4. form ionic bonds with water molecules

7. When discussing solubility of sulfur dioxide in water, all of the following need to be considered except
   1. pressure of the gas
2. temperature
3. nature of the solvent
4. molecular mass of solute

8. Which of the following salts is most soluble at 60°C?
   1. KNO₃
   2. NaNO₃
   3. KCl
   4. NaCl

9. The solubility of a gas increases as
   1. the temperature increases
   2. the pressure increases
   3. the pressure decreases
   4. the solvent becomes more polar

10. Temperature has the least effect on the solubility of
    1. NaCl
    2. KCl
    3. HCl
    4. NH₄Cl

True/False:

11. _____ The solubility of HCl increases as the temperature decreases.
12. _____ Salt in water is an example of a solid in liquid solution.
13. _____ Decreasing surface area enhances solute contact with solvent.
14. _____ As the partial pressure increases, the solubility of a gas increases.
15. _____ Gas solubilities are commonly reported in moles/L.
16. _____ Carbon dioxide in ether is an example of a gas in liquid solution.
17. _____ Increasing collisions of solvent with solute increases the rate of solution formation.
18. _____ At 20°C, KCl is more soluble than NaCl.
19. _____ The kinetic energy of a gas increases as the temperature increases.
20. _____ Solid-solid solutions are called alloys.

Fill in the Blank:

21. Define the following terms
   1. solubility
   2. Henry’s law
   3. saturated solution
   4. recrystallization
   5. solution equilibrium

22. The solubility of a gaseous solute is affected by both the ______ and the ______ of the gas.
23. There are approximately ________ g NaNO₃ dissolved per 100 g water at 40°C.
24. Dissolving of a solid by water depends upon the ______ that occur between the ______ molecules and the particles in the solid ______.

Short Answers:

25. The solubility of a certain gas is 0.85 g/L at 20°C and 650 mm Hg. What is the solubility when the gas pressure is increased to 1340 mm Hg?
26. The solubility of a gas is 1.25 g/L at 920 mm Hg. At what pressure will the gas solubility be 0.92 g/L?
Answer Key

1. c
2. b
3. a
4. b
5. a
6. b
7. d
8. a
9. b
10. a

11. true
12. true
13. false
14. true
15. false
16. true
17. true
18. false
19. true
20. true

1. the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature.
2. the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid.
3. a solution that contains the maximum amount of solute that is capable of being dissolved.
4. the process of dissolved solute returning to the solid state.
5. the physical state described by the opposing processes of dissolution and recrystallization occurring at the same rate.

21. temperature, pressure
22. 105
23. collisions, solvent, crystals
24. \( S_2 = \frac{S_1 \times P_2}{P_1} = \frac{0.85 \text{ g/L} \times 1340 \text{ mm Hg}}{650 \text{ mm Hg}} = 1.75 \text{ g/L} \)
25. \( P_2 = \frac{S_2 \times P_1}{S_1} = \frac{0.92 \text{ g/L} \times 920 \text{ mm Hg}}{1.25 \text{ g/L}} = 677.1 \text{ mm Hg} \)
16.2 Solution Concentration

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. The symbol for molality is
   1. m
   2. \( m \)
   3. M
   4. \( M \)

2. The symbol for molarity is
   1. m
   2. \( m \)
   3. M
   4. \( M \)

3. Mass percent is determined by the calculation
   1. \( \frac{\text{grams solute}}{\text{grams solvent}} \times 100\% \)
   2. \( \frac{\text{grams solute}}{\text{grams solution}} \times 100\% \)
   3. \( \frac{\text{grams solvent}}{\text{grams solute}} \times 100\% \)
   4. \( \frac{\text{grams solute} \times \text{grams solvent}}{\text{grams solvent}} \times 100\% \)

4. In making a volume % solution, the volume of solute
   1. depends on the solvent type
   2. is calculated as \( \frac{\text{solute}}{\text{solvent}} \times \text{final volume} \)
   3. is added to the desired final volume
   4. is subtracted from the desired final volume

5. One of the following concentration units allows us to determine the number of particles present.
   1. volume percent
   2. percent dilution
   3. molarity
   4. mass percent

6. Units for molarity are
   1. grams/liter
   2. grams/mL
   3. moles %
   4. moles/L

7. To determine molarity, we need to know _______ of solute
   1. grams/L
   2. grams%
16.2. Solution Concentration

3. grams/mole
4. moles/gram

8. Molality calculations require that we know
   1. liters of solution
   2. liters of solvent
   3. kilograms of solvent
   4. kilograms of solute

9. The best glassware to use in preparing a solution is a
   1. graduated cylinder
   2. volumetric flask
   3. calibrated beaker
   4. buret

10. The formula to use for calculations of dilutions is
    1. \( V_1C_1 = V_2C_2 \)
    2. \( V_1C_2 = V_2C_1 \)
    3. \( \frac{V_1}{C_1} = \frac{V_2}{C_2} \)
    4. \( \frac{C_1}{V_1} = \frac{C_2}{V_2} \)

True/False:

11. _____ Molarity values are not affected by temperature
12. _____ The concentration of a solution is decreased by adding water to it.
13. _____ You can use a 1.0 liter flask to prepare 500 mL of solution.
14. _____ A dilute solution contains a relatively small amount of solute
15. _____ If the solvent evaporates from a solution, that solution becomes less concentrated.
16. _____ Mass percent is used if the solute is a solid.
17. _____ Volume percent is calculated using volume solute ÷ volume solvent.
18. _____ Volume is not preserved in many liquid solute/solvent combinations.
19. _____ “Concentrated” and “dilute” provide accurate information about solute concentrations.
20. _____ It is not important to know whether a % solution is by mass or by volume.

Short Answers:

21. Describe how to prepare 250 mL of a 5% (mass/mass) solution of glucose in water.
22. Describe how to prepare 500 mL of a 10% (vol/vol) solution of acetone in water.
23. How many mL of 0.45 M NaCl solution will you need to prepare 150 mL of 0.1 M NaCl?
24. You dilute 200 mL of a glucose stock solution to make 500 mL of a 0.30 M solution. What was the concentration of the stock solution?
25. Describe how to prepare 500 mL of 0.2 M LiCl.
26. You dissolve 45 g methanol (molar mass = 32 g/mole) in water and dilute to 300 mL. What is the molarity of the solution?
27. How many grams of MgCl₂ would you dissolve in 1.5 kilograms water to form a 0.5 m solution?
28. What is the molality of a solution in which 0.77 moles of solute are dissolved in 0.45 kg water?
29. Why would it be difficult to prepare a 0.5 M solution of CO₂?
30. You are preparing molal solutions by dissolving 0.15 moles of solute into 500 grams of solvent. Solvent A has a density of 0.6 g/mL and solvent B has a density of 1.4 g/mL. Which solution will contain the higher concentration of solute?
**Answer Key**

1. b
2. c
3. b
4. d
5. c
6. d
7. c
8. c
9. b
10. a
11. false
12. true
13. false
14. true
15. false
16. true
17. false
18. true
19. false
20. false
21. 

\[
\text{mass of solute} = \left( \frac{\text{percent by mass}}{100\%} \right) \times \text{mass of solution}
\]

\[
= \frac{5\%}{100\%} \times 250 \text{ g}
\]

\[
= 12.5 \text{ g glucose}
\]

22. Dissolve 12.5 g glucose in (250-12.5) g water or 237.5 mL water.

23. volume solute = (10% by \(\frac{\text{vol}}{100\%}\)) \times 500 mL = 50 mL

24. dissolve 50 mL acetone in water and dilute to 500 mL.

25. \(V_1C_1 = V_2C_2\) so \(V_1 = \frac{V_2C_2}{C_1} = V_1 = \frac{150 \text{ mL} \times 0.1 \text{ M}}{0.45 \text{ M}} = 33.3 \text{ mL}\)

26. \(V_1C_1 = V_2C_2\) so \(C_1 = \frac{V_2C_2}{V_1} = C_1 = \frac{500 \text{ mL} \times 0.3 \text{ M}}{200 \text{ mL}} = 0.75 \text{ M}\)

27. molar mass LiCl = 42.4 g/mol

28. 42.4 g/mol \times 0.2 \text{ mol/L} \times 0.500 L = 4.24 g LiCl

29. Dissolve in about 200 mL, mix well, and dilute to 500 mL.

30. 45 g ÷ 32 g/mol = 1.41 mol \(\frac{1.41 \text{ mol}}{0.3 \text{ L}} = 4.69 \text{ M}\)

31. molar mass MgCl₂ is 95.3 g/mole need 0.5 moles in 1.5 kg

32. 0.5 \text{ mol} = \text{ moles MgCl}_2 \div 1.5 \text{ kg} = 0.5 \text{ moles/kg} \times 1.5 \text{ kg} = 0.75 \text{ moles MgCl}_2

33. 0.75 \text{ moles} \times 95.3 \text{ g/mol} = 71.5 \text{ g MgCl}_2 \text{ needed}

34. molality = \(\frac{\text{moles solute}}{\text{kg solvent}} = \frac{0.77 \text{ moles}}{0.45 \text{ kg}} = 1.71 \text{ molal}\)

35. There is no easy way to weigh out a specific mass of CO₂. In addition, the amount dissolved depends on the temperature and the pressure.

36. Solvent A volume used: 500 g ÷ 0.6 g/mL = 833.3 mL so \(\frac{0.15 \text{ moles}}{0.833 \text{ L}} = 0.18 \text{ M solution}\)

37. Solvent B volume used:

38. 500 g ÷ 1.4 g/mL = 357.1 mL so \(\frac{0.15 \text{ moles}}{0.357 \text{ L}} = 0.42 \text{ M solution}\)

39. Solvent B has the higher concentration of solute.
16.3 Colligative Properties

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. A colligative property is one based on
   1. number of covalent bonds
   2. number of hydrogen bonds
   3. number of particles
   4. number of dipole-dipole interactions

2. The freezing point of a solution is lowered by
   1. increasing the number of solute particles
   2. increasing the kinetic energy
   3. decreasing the number of solute particles
   4. increasing the pressure on the solution

3. The freezing point of water will be lowered most by dissolving 1.0 mole of
   1. naphthalene
   2. NaCl
   3. MgCl₂
   4. ether

4. The best solute for lowering vapor pressure is one that
   1. is slightly soluble
   2. evaporates readily
   3. is nonvolatile
   4. reacts with the solvent

5. Solvent freezing points are lowered by materials that
   1. bring more order to the solvent molecules
   2. disrupt the order of solvent molecules
   3. increase kinetic energy of solvent molecules
   4. decrease the kinetic energy of solvent molecules

6. Vapor pressure is a ______ property of the solvent
   1. physical
   2. chemical
   3. dynamic
   4. kinetic

7. $K_f$ stands for
   1. freezing point elevation constant
   2. molar freezing point constant
3. micro freezing point constant
4. molal freezing point constant

8. A colligative property of water is
   1. molecular structure
   2. boiling point
   3. solubility
   4. reactivity

9. A volatile liquid
   1. evaporates readily
   2. has a low vapor pressure
   3. has a high boiling point
   4. is unreactive

10. Molality is defined as
    1. moles solute/moles solvent
    2. moles solute/moles solution
    3. moles solute/kg solution
    4. moles solute/kg solvent

True/False:

11. _____ Ethylene glycol is used in many commercial antifreeze preparations.
12. _____ NaCl will depress the freezing point of water more than CaCl₂.
13. _____ A colligative property is one that depends on the identity of particles.
14. _____ Liquids that do not evaporate easily have lower vapor pressures.
15. _____ Solutes used to lower vapor pressure cannot evaporate readily.
16. _____ The amount of freezing point depression is inversely proportional to the number of particles dissolved.
17. _____ Vapor pressure is determined by how rapidly molecules can escape the surface of a liquid.
18. _____ When a solute is added to a solvent, more energy must be removed from the solution in order to freeze it.
19. _____ Salt is added to an icy road to increase the boiling point of the water.
20. _____ The vapor pressure of a solution is lower, so less energy is needed to bring that pressure up to the pressure of the atmosphere.

Fill in the Blank:

21. Define the following terms:
    1. freezing point depression
    2. molal freezing point depression constant
    3. boiling point elevation
    4. molal boiling point elevation constant

22. When a pure _____ freezes, its particles become more _____ as the ______ forces that operate between the molecules fix each molecule in place.

23. While _____ does not depend on temperature, the _____ of the solution increases with temperature.

Short Answers:

24. We add 320 g glucose (molar mass 180 g/mole) to 0.65 kg water. What will the new freezing point be?
25. A solution is prepared of 60 g KCl in 0.700 kg water. What is the boiling point of the new solution?
26. A solution containing 15.00 g of a carbohydrate in 100.0 g of water is found to freeze at -1.53°C. What is the molar mass of he carbohydrate?
16.3. Colligative Properties

Answer Key

1. c
2. a
3. c
4. c
5. b
6. a
7. d
8. b
9. a
10. d
11. true
12. false
13. false
14. true
15. true
16. false
17. true
18. true
19. false
20. false

1. the difference in temperature between the freezing point of a pure solvent and that of a solution.
2. a constant that is equal to the change in the freezing point for a 1-molal solution of a nonvolatile molecular solute.
3. the difference in temperature between the boiling point of the pure solvent and that of the solution.
4. a constant that is equal to the change in the boiling point for a 1-molal solution of a nonvolatile molecular solute.

21. solvent, ordered, intermolecular
22. mass, volume
23. 320 g ÷ 180 g/mole = 1.78 moles
24. 1.78 moles ÷ 0.65 kg = 2.74 m
25. \( \Delta T_f = K_f \times m = -1.86^\circ C/m \times 1.79 m = -3.32^\circ C \)
26. the new freezing point for the solution
27. KCl dissociates into two particles/molecule
28. 60 g ÷ 74.6 g/mole = 0.80 moles
29. 0.80 moles ÷ 0.7 kg = 1.14 m
30. \( \Delta T_b = K_b \times m = 0.512^\circ C/m \times 1.14 m \times 2 = 1.16^\circ C \)
31. The new boiling point = 101.16°C
32. \( m = \Delta T_f \div K_f = -1.53^\circ C \div -1.86^\circ C/m = 0.82 m \)
33. mol solute = 0.82 m × 0.1 kg = 0.082 moles
34. 15 g ÷ 0.082 moles = 182.9 g/mole
16.4 Net Ionic Equations

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. The symbol \((aq)\) represents
   1. water molecules as part of the reaction
   2. hydrated particles
   3. water of hydration
   4. water as a product

2. A net ionic equation shows
   1. all the ionic components in the reaction
   2. spectator ions
   3. ions that are directly involved
   4. both ions and molecules

3. Spectator ions
   1. help balance net ionic equations
   2. form products
   3. balance charges
   4. are not involved in the reaction

4. The compound \(\text{CaSiO}_3\) is
   1. mostly insoluble
   2. mostly soluble
   3. highly soluble
   4. soluble

5. A single-replacement reaction involves all of the following as products except
   1. metal
   2. water
   3. hydrogen
   4. halogen

6. One of the following would not be a reactant in a single-replacement reaction
   1. Zn
   2. HCl
   3. \(\text{CH}_3\text{OH}\)
   4. \(\text{I}_2\)

7. All of the following are mostly soluble except
   1. CsI
   2. NaBr
3. LiCl
4. AgBr

8. All of the following are mostly insoluble except
   1. NaOH
   2. Ca(OH)\(_2\)
   3. Mg(OH)\(_2\)
   4. Zn(OH)\(_2\)

9. All of the following are soluble except
   1. (NH\(_4\))\(_2\)CO\(_3\)
   2. PbCO\(_3\)
   3. NaCl
   4. Pb(NO\(_3\))\(_2\)

10. Only one of the following is soluble
    1. KCl
    2. AgCl
    3. PbCl\(_2\)
    4. Hg\(_2\)Cl\(_2\)

True/False:

11. _____ Silver nitrate is completely soluble in water.
12. _____ Sodium acetate is only somewhat soluble in water.
13. _____ One reactant in a single replacement reaction is a pure neutral element.
14. _____ Single-replacement reactions do not involve spectator ions.
15. _____ In an ionic reaction, only the ions of interest are shown.
16. _____ A gas and water can be products of a double-replacement reaction.
17. _____ If copper(II) phosphate is formed in a reaction, it would be designated as (aq).
18. _____ The designation (s) indicates a material in solution.
19. _____ A single-replacement reaction might have water as a reactant.
20. _____ Cesium sulfide is a soluble material.

Short Answers:

21. Write the molecular equation, the ionic equation and the net ionic equation for the following double replacement reactions:
    1. silver nitrate and rubidium chloride  
    2. mercury(I) nitrate and hydrochloric acid  
    3. calcium chloride and sodium carbonate

22. Complete the following equations, balance the equations, and indicate any insoluble products with (s)
    1. Al + CuCl\(_2\) \(\rightarrow\)
    2. Br\(_2\) + CaI\(_2\) \(\rightarrow\)
    3. Al + HCl \(\rightarrow\)
    4. Mg + HCl \(\rightarrow\)
    5. Zn + H\(_2\)SO\(_4\) \(\rightarrow\)

Answer Key

1. b
2. c
3. d
4. a
5. b
6. c
7. d
8. a
9. b
10. a
11. true
12. false
13. true
14. true
15. false
16. true
17. false
18. false
19. true
20. true
16.5 Solutions

Chapter Test

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. The solubility of a gas decreases as
   1. the pressure increases
   2. the temperature increases
   3. the temperature decreases
   4. the solvent is less polar

2. A solid solute contact with solvent is enhanced by all of the following except
   1. increasing surface area
   2. stirring
   3. crushing the solute
   4. increasing pressure

3. The role of water in aqueous solution formation is to
   1. hydrate solute particles
   2. form ionic bonds with solute
   3. be surrounded by solute molecules
   4. react with solute molecules

4. Molarity allows us to calculate directly the following
   1. freezing point depression
   2. number of particles in solution
   3. percent hydration
   4. partial pressure

5. One of the following is not affected by temperature
   1. molarity values
   2. gas solubility
   3. molality values
   4. solid solute solubility

6. Boiling point is a _________ property of water
   1. collective
   2. collaborative
   3. colligative
   4. collegial

7. The boiling point of a liquid is raised by
   1. increasing the number of solute particles
   2. increasing the pressure
3. decreasing the number of solute particles
4. using nonelectrolytes instead of electrolytes

8. The symbol (s) stands for
   1. solid
   2. solute
   3. solvent
   4. solution

9. All alkali metal salts are
   1. insoluble
   2. mostly insoluble
   3. mostly soluble
   4. soluble

10. Most compounds containing the silicate ion are
    1. insoluble
    2. mostly insoluble
    3. mostly soluble
    4. soluble

True/False:

11. _____ The symbol for molality is m.
12. _____ Units for molarity are moles/liter.
13. _____ Dilution calculations can be made using \( \frac{V_1}{C_1} = \frac{V_2}{C_2} \).
14. _____ Alloys are solutions of solids in other solids.
15. _____ Oxygen in water is an example of a gas in liquid solution.
16. _____ The freezing point of a liquid will be lowered when solute particles are removed.
17. _____ Vapor pressure is a physical property of a solution.
18. _____ The compound magnesium chromate is very soluble in water.
19. _____ Net ionic equations include spectator ions.
20. _____ A halogen is one possible product of a single-replacement reaction.

Short Answers:

21. Define the following terms:
   1. recrystallization
   2. boiling point elevation
   3. saturated solution
   4. solution equilibrium

   1. The solubility of a gas in water is 0.56 g/L at 1.00 atm. What will be the solubility when the pressure is raised 3.9 atm?
   2. The solubility of a gas is 2.5 g/L at 3560 mm Hg. What pressure will the solubility be 1.7 g/L?

22. Describe how to prepare 250 mL of a 12% solution of potassium bromide.
23. You dissolve 35 g of Mg(NO₃)₂ in water and dilute to 1.0 L. What is the molarity of this solution?
24. How would you prepare 100 mL of a 0.05 m solution of CsCl in water?
25. You dissolve 139 g ethylene glycol (molar mass = 62.07 g/mole) in 500 g water. How much will that solution lower the freezing point of water?
26. Write the molecular equation, the ionic equation and the net ionic equation for each of the following reactions:
   1. lithium sulfide and nickel(II) nitrate
2. barium chloride and sodium sulfate

27. Write balanced equations for the following reactions:
   1. aluminum + zinc oxide
   2. lithium + calcium sulfide.

---

**Answer Key**

1. b
2. d
3. a
4. b
5. c
6. c
7. a
8. a
9. d
10. b
11. false
12. true
13. false
14. true
15. true
16. false
17. true
18. false
19. false
20. true

1. the process of dissolved solute returning to the solid state.
2. the difference in temperature between the boiling point of the pure solvent and that of the solution.
3. a solution that contains the maximum amount of solute that is capable of being dissolved.
4. the physical state described by the opposing processes of dissolution and recrystallization occurring at the same rate.

23. \[ \% \text{solution} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% \]
24. mass of solute = (\%solution \times mass of solution) \div 100\%
25. \( 0.12 \times 250 \, g = 30 \, g \, KBr \)
26. mix 30 g KBr with 220 grams (or mL) of water.
27. \( \text{Mg(NO}_3)_2 \) molar mass = 148.3 g/mole
28. \( 35 \, g \div 148.3 \, g/mole = 0.23 \, moles \) dissolved in \( 1.0 \, L = 0.23 \, moles/l = 0.23 \, M \)
29. \( \text{CsCl} \) molar mass = 168.4 g/mole
30. \( 0.05 \, moles \times 168.4 \, g/mole = 8.42 \, g \) sow would want 8.42 g/kg water
31. need \( 100 \, mL \) or 0.1 kg so need \( 0.1 \times 8.42 = 0.842 \, grams \) \( \text{CsCl} \) added to \( 100 \, mL \) water.
32. \( 139 \, g \div 62.07 \, g/mole = 2.24 \, moles \)
33. \( 2.24 \, moles \div 0.5 \, kg \) water = 4.48 m
34. \( K_f = -1.86^\circ C/m \) so \( 4.48 \, m \times -1.86^\circ C/m = -8.33^\circ C \) new freezing point

   a. molecular equation – \( \text{Li}_2\text{S(aq)} + \text{Ni(NO}_3)_2(aq) \rightarrow 2\text{LiNO}_3(aq) + \text{NiS(s)} \)
   b. ionic equation –
c. $2Li^+(aq) + S^{2-}(aq) + Ni^{2+}(aq) + 2NO_3^-(aq) \rightarrow 2Li^+(aq) + 2NO_3(aq) + NiS(s)$

d. net ionic equation –

e. $Ni^{2+}(aq) + S^{2-}(aq) \rightarrow NiS(s)$

f. molecular equation $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

g. ionic equation –

h. $Ba^{2+}(aq) + 2Cl^-(aq) + 2Na^+(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) + 2Na^+(aq) + 2Cl^-(aq)$

i. net ionic equation –

j. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

1. $2Al + 3ZnO \rightarrow Al_2O_3 + 3Zn$

2. $Li + CaS \rightarrow Li_2S + Ca$
Chapter 17
Thermochemistry
Assessments

Chapter Outline

17.1 Heat Flow
17.2 Thermochemical Equations
17.3 Heat and Changes of State
17.4 Hess’s Law
17.5 Thermochemistry
17.1 Heat Flow

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. In order to be useful, chemical potential energy must be
   1. stored
   2. sequestered
   3. released
   4. organized

2. All of the following statements about heat are true except
   1. heat is a form of energy
   2. heat is transferred from one place to another
   3. chemical reactions can produce heat
   4. heat energy moves from a lower temperature to a higher temperature

3. A system is
   1. the matter in a given space not involved in a reaction
   2. matter in a given space that is involved in the reaction
   3. matter in a given space that is not acted upon
   4. matter in a given space that has already reacted

4. If heat is absorbed from the surroundings, the process is
   1. endothermic
   2. exothermic
   3. endodynamic
   4. exodynamic

5. One of the following is a unit of heat
   1. caloris
   2. joule
   3. caloric
   4. jole

6. After one hour exposure to the sun, which of the following will have the greatest increase in temperature?
   1. shallow dish of water
   2. gallon bucket of water
   3. swimming pool
   4. lake

7. Units for specific heat are
   1. joules/g
   2. joules/kg
3. joules/g·°C
4. joules/kg·°C

8. Water is used as ______ in power plants
   1. power source
   2. fuel
   3. coolant
   4. thermostat

9. The m in the specific heat formula stands for
   1. mass in grams
   2. moles
   3. mass in kg
   4. molal

10. The symbol for specific heat in the equation is
   1. $e_p$
   2. $p_c$
   3. $h_p$
   4. $c_p$

**True/False:**

11. _____ A Calorie is the equivalent of 500 calories.
12. _____ When set in the sun at the same time, a gold coin will heat more rapidly than a silver coin.
13. _____ A fireplace warms a room because the heat produced by the fire is absorbed by air particles.
14. _____ Brewing coffee is an exothermic process.
15. _____ The sign of $q$ is positive for an endothermic process.
16. _____ A calorie is the amount of energy needed to raise the temperature of one gram of water by one degree C.
17. _____ One joule equals 4.184 calories.
18. _____ Water has a high specific heat.
19. _____ When two objects at different temperatures come in contact, heat flows from the lower temperature object to the higher temperature object.
20. _____ Metals are not very resistant to changes in temperature.

**Fill in the Blank:**

21. The _____ changes of a system occur as either _____ or work, or some combination of both.
22. Heat is _____ that is transferred from one object or substance to another because of a difference in ______-____ between them.
23. ______ is the study of energy changes that occur during ______ reactions and during changes of ______.
24. The ______ is the specific portion of matter in a given space that is being studied during an ______ or an ______.

**Short Answers:**

25. How many calories are there in 504 joules?
26. A 500 g sample of water loses heat as the temperature drops from 29.4°C to 19.7°C. How much energy was lost?
27. A sample of gold is heated from 12°C to 28°C and absorbs 460 joules of energy. What is the mass of the gold?
28. A 25.2 g sample of tungsten absorbs 734 J of heat. The temperature increases from 14.5°C to 85.4°C. What is the specific heat of tungsten?
Answer Key

1. c
2. d
3. b
4. a
5. b
6. a
7. c
8. c
9. a
10. d
11. false
12. true
13. true
14. false
15. true
16. true
17. false
18. true
19. false
20. false
21. energy, heat
22. energy, temperature
23. Thermochemistry, chemical, state
24. system, experiment, observation
25. 504 joules × \frac{calorie}{4184 joules} = 120.5 calories
26. heat = m \times c_p \times \Delta T = 500 g \times 4.18 J/g°C \times 9.7°C = 20273 J = 20.27 kJ
27. heat = m \times c_p \times \Delta T \text{ so } m = \frac{\text{heat}}{c_p \times \Delta T} = \frac{460 J}{4.18 J/g°C \times 16°C} = 6.9 g
28. heat = m \times c_p \times \Delta T \text{ so } c_p = \frac{\text{heat}}{m \times \Delta T} = \frac{236 J}{25.2 g \times 10.9°C} = 0.134 J/g°C
Lesson Quiz

Multiple Choice:

1. Enthalpy changes are commonly measured at
   1. constant temperature
   2. constant pressure
   3. constant volume
   4. constant concentration

2. $\Delta H$ is the designation for
   1. enthalpy change
   2. entropy change
   3. entropic change
   4. enthalpic change

3. The designation for heat is
   1. h
   2. e
   3. q
   4. m

4. Foam cups are commonly used to make an inexpensive
   1. coulomb meter
   2. colorimeter
   3. caloric meter
   4. calorimeter

5. Studies of heat change would be easiest for the following type of reaction:
   1. single-replacement
   2. formation of a gas
   3. combustion
   4. formation of a precipitate

6. One of the following is not a characteristic of an endothermic reaction
   1. heat goes from surroundings to system
   2. heat involved is written on the reactant side of the equation
   3. heat is released
   4. energy is in kJ

7. Studies of enthalpy involve measuring
   1. change of state
   2. temperature change
3. gas formation
4. solid formation

8. One key to successful measurements of heat of reaction is
   1. no temperature exchange with container and room
   2. reaction must form a gas
   3. precipitate formation does not interfere with measurement
   4. temperature must increase

9. A negative enthalpy changes means that
   1. heat is absorbed from the surroundings
   2. energy must be put into the reaction
   3. heat is given off to the surroundings
   4. the energy change in the reaction is negligible

10. The heat of a reaction is equivalent to
    1. $\Delta H$
    2. $\Delta G$
    3. $\Delta G^\circ$
    4. $\Delta H^r$

True/False:

11. _____ Chemical reactions in the lab are run with controlled temperature.
12. _____ Enthalpy measurements are made only when heat is released.
13. _____ A temperature change occurs when reactants are converted to products.
14. _____ A lid is used to control pressure in the reaction system
15. _____ The dissolved materials are the system.
16. _____ It is important to measure solution temperatures both before and after a reaction has occurred.
17. _____ The enthalpy change for an endothermic reaction is negative.
18. _____ The value of $\Delta H$ depends on the physical state of reactants and products.
19. _____ Energy changes follow stoichiometric relationships.
20. _____ Heats of reaction are usually measured in kJ.

Fill in the Blank:

21. _______ is the heat content of a _______ at constant _______.
22. _______ is the measurement of the transfer of _______ into or out of a system during a _______ reaction or _______ process.

Short Answers:

23. In the reaction $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g); \Delta H = -196.4 \text{ kJ}$. calculate the energy change when 1.0 g hydrogen peroxide decomposes.
24. When 15.3 g of sodium nitrate, NaNO$_3$, was dissolved in water in a calorimeter, the temperature fell from 25.00°C to 21.56°C. What is the enthalpy change when one mole of NaNO$_3$ dissolves?
25. A piece of metal is heated and dropped into a foam cup calorimeter containing 145 mL of water at 20.5°C. The temperature of the water rises to 26.7°C. How many joules of heat were released?
26. What is the specific heat of a substance that absorbs $2.5 \times 10^3$ joules of heat when a sample of $1.0 \times 10^4$ g of the substance increases in temperature from 10.0 C to 70.0 C?
**Answer Key**

1. b  
2. a  
3. c  
4. d  
5. c  
6. c  
7. b  
8. a  
9. c  
10. a  
11. false  
12. false  
13. true  
14. false  
15. true  
16. true  
17. false  
18. true  
19. true  
20. true  

21. enthalpy, system, pressure  
22. calorimetry, heat, chemical, physical  
23. molar mass of $H_2O_2 = 34.0 \text{ g/mole}$  
24. two moles hydrogen peroxide produce -196.4 kJ, or 98.2 kJ/mol  
25. $\Delta H = 1.0 \text{ g } H_2O_2 \times \frac{1\text{ mol}}{34\text{ g}} \times \frac{-98.2 \text{ kJ}}{\text{ mol}} = 2.89 \text{ kJ}$  
26. $\Delta H = m \times c_p \times \Delta T = 15.3 \text{ g } \times \frac{4.18 \text{ J}}{\text{ g}^\circ \text{C}} \times 3.44^\circ = 220 \text{ J}$  
27. molar mass  
28. $NaNO_3 = 85 \text{ g/mol}$  
29. so  
30. $\frac{15.3 \text{ g}}{85 \text{ g/mol}} = 0.18 \text{ mol}$  
31. $220 \text{ J} \div 0.18 \text{ mol} = 1222 \text{ J/mol} = 1.22 \text{ kJ/mol}$  
32. $\Delta H = m \times c_p \times \Delta T = 145 \text{ g } \times \frac{4.18 \text{ J}}{\text{ g}^\circ \text{C}} \times 6.2^\circ \text{C} = 3758 \text{ joules } = 3.758 \text{ kJ}$  
33. $\Delta H = m \times c_p \times \Delta T$ so $c_p = \frac{\Delta H}{(m \times \Delta T)} = \frac{2500 \text{ J}}{1 \times 10^5 \text{ g} \times 60^\circ \text{C}} = 0.0042 \text{ J/}^\circ \text{C}$
Lesson Quiz

Multiple Choice:

1. The heat of solidification for water is
   1. 6.01 kJ/mol
   2. -6.01 kJ/mol
   3. 6.01 J/mol
   4. -6.01 J/mol

2. The heat of solidification is strongly influenced by
   1. temperature
   2. pressure
   3. intermolecular forces
   4. solvent

3. The heat of solidification deals with
   1. liquid converting to solid
   2. solid converting to gas
   3. solid converting to liquid
   4. liquid converting to gas

4. As steam condenses to liquid water, the temperature of the steam
   1. increases
   2. decreases
   3. does not change
   4. fluctuates

5. Changes from the vapor state to the liquid state involves all of the following except
   1. release of energy
   2. less space between the particles in the liquid state
   3. disruption of intermolecular forces
   4. particles closer together in the liquid state

6. Ethanol has a higher heat of vaporization than other materials in the table because
   1. ethanol has a higher molecular mass
   2. ethanol can form more hydrogen bonds
   3. ethanol contains more hydrogen atoms
   4. ethanol contains more carbons

7. Calculation of the energy needed to melt 156 grams of ice uses the
   1. molar heat of fusion
   2. specific heat of water
3. molar heat of vaporization  
4. specific heat of steam

8. The process of converting water at 55°C to ice at -12°C requires _____ steps.
   1. five  
   2. four  
   3. three  
   4. two

9. The molar heat of solution for KClO₄ is 51.04 kJ/mol. This value tells us that KClO₄ will
   1. dissolve slowly in water  
   2. generate a great deal of heat when dissolving  
   3. increase the specific heat of the solution  
   4. decrease solvent temperature while dissolving

10. Many cold packs use
   1. sodium hydroxide  
   2. ammonium nitrate  
   3. calcium chloride  
   4. sodium chloride

**True/False:**

11. _____ The conversion of solid directly to vapor is known as sublimation.
12. _____ Energy is released when steam condenses to liquid water.
13. _____ The total energy released or absorbed in a process is independent of the amount of material involved.
14. _____ Energy is absorbed when a gas is converted to a liquid.
15. _____ The molar heat of vaporization is used to determine the amount of energy needed to convert water to ice.
16. _____ Calcium chloride releases heat energy when it dissolves.
17. _____ A temperature drop will be experienced when a solution of NaOH is prepared.
18. _____ For all substances, the heat of vaporization is greater than the heat of fusion.
19. _____ The molar heat of vaporization and the molar heat of condensation have different numerical values.
20. _____ units for molar heat of fusion are kJ/gram.

**Fill in the Blank:**

21. The molar heat of fusion of a substance is the heat _______ by one mole of that substance as it is converted from a _______ to a liquid.
22. Every substance has a _______ value for its molar heat of fusion, depending on the amount of energy required to disrupt the _______ forces present in the solid.
23. The molar heat of condensation of a substance is the heat _______ by one mole of that substance as it is converted from a _______ to a liquid.
24. The molar heat of _________a substance is the heat _______ or _______ when one _______ of the substance is dissolved in water.

**Short Answers:**

25. How much heat is absorbed when 43.9 grams of ethanol are converted from the liquid to the vapor state?
26. How much heat is released when 125 g steam at 115°C are converted to liquid water at 85°C?
27. How many grams of methanol are present if the conversion of liquid to solid methanol releases 16.43 kJ of energy?
28. The molar heat of solution for KI is 20.33 kJ/mol. What will be the temperature change if 47 g KI are dissolved in 1500 mL water at 20°C?
Answer Key

1. b
2. c
3. a
4. c
5. c
6. a
7. a
8. c
9. d
10. b
11. true
12. true
13. false
14. false
15. false
16. true
17. false
18. true
19. false
20. false

21. absorbed, solid
22. unique, intermolecular
23. released, gas
24. solution, absorbed, released, mole
25. molar mass of ethanol is 46 g/mole
26. moles ethanol = \( \frac{43.9 \text{ g}}{46 \text{ g/mole}} \) = 0.95 moles
27. 43.9 kJ/mol \times 0.95 \text{ moles} = 41.3 kJ

28. molar mass of water = 18.01 grams/mole we have \( \frac{125 \text{ g}}{18.01 \text{ g/mole}} = 6.94 \text{ moles} \)

a. steam at 115°C to steam at 100°C
b. \( \Delta H = m \times c_p \times \Delta T = 125 \text{ g} \times \frac{1.87 \text{ J}}{\text{g} \cdot ^\circ \text{C}} \times 15^\circ \text{C} = 3506 \text{ J} = 3.5 \text{ kJ} \)
c. steam at 100°C to liquid at 100°C
d. 6.94 moles \times 40.7 \text{ kJ/mol} = 282.5 \text{ kJ}
e. water at 100°C to water at 85°C
f. 125 \text{ g} \times \frac{4.18 \text{ J}}{\text{g} \cdot ^\circ \text{C}} \times 15^\circ \text{C} = 7838 \text{ J} = 7.38 \text{ kJ}
g. 3.5 \text{ kJ} + 282.5 \text{ kJ} + 7.38 \text{ kJ} = 293.4 \text{ kJ}

27. molar mass of methanol = 32 g/mole
28. \( \frac{16.43 \text{ kJ}}{\text{mole}} \times \frac{15.16 \text{ g}}{\text{mol}} \times \frac{32 \text{ g}}{\text{mole}} = 166.4 \text{ g} \)
29. \( 47 \text{ g} \times \frac{1000 \text{ J}}{166 \text{ g}} \times \frac{20.33 \text{ kJ}}{\text{mol}} \times \frac{1 \text{ mol}}{1000 \text{ J}} = 5756 \text{ J} \)
30. \( \Delta T = \frac{q}{c_p \times m} = \frac{5756 \text{ J}}{4.18 \text{ J/g} \cdot ^\circ \text{C} \times 1547 \text{ g}} = 0.9^\circ \text{C} \)
31. specific heat value was positive, so temperature would drop 0.9°C to 19.1°C
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. Hess’ law deals with
   1. adding thermochemical equations
   2. calculating change of state situations
   3. predicting products of reactions
   4. energy transitions

2. Calorimetry is a ______ way to measure heats of reaction.
   1. calculated
   2. theoretical
   3. direct
   4. indirect

3. Heats of combustion in a table are for
   1. combustion of one gram of substance
   2. formation of one kJ of energy
   3. utilization of one kJ of energy
   4. combustion of one mole of substance

4. If the coefficients of a reaction ae doubled, the $\Delta H$ values are
   1. unchanged
   2. doubled
   3. divided by two
   4. squared

5. $\Delta H^\circ$ is the symbol for
   1. standard heat of reaction
   2. standard heat of formation
   3. standard heat of molar formation
   4. standard heat of combustion

6. $\Delta H^r$ is the symbol for the
   1. enthalpy change under experimental conditions
   2. entropy change under experimental conditions
   3. enthalpy change under standard conditions
   4. enthalpy change under standard conditions

7. The symbol $n$ in the standard heat of reaction equation stands for
   1. number of chemicals in the reaction
   2. coefficient of energy
3. number of molecules in product
4. coefficient of a chemical in the reaction

8. $\Delta H^\circ$ is called
   1. standard enthalpy of reaction
   2. standard enthalpy of change
   3. standard enthalpy
   4. standard enthalpy of product

9. The standard state for bromine is
   1. gas at 25°C and 101.3 kPa
   2. gas at 20°C and 101.3 kPa
   3. liquid at 25°C and 101.3 kPa
   4. liquid at 20°C and 101.3 kPa

10. A theoretical reaction for the synthesis of acetylene involves a reaction between
    1. graphite and air
    2. graphite and hydrogen gas
    3. hydrogen gas and methane
    4. hydrogen gas and ethene

**True/False:**

11. _____ Adding heats of reaction is an indirect approach to determining enthalpy change.
12. _____ Enthalpy changes for combustion reactions are difficult to measure.
13. _____ When coefficients in reactions are doubled, $\Delta H$ does not change.
14. _____ The standard state for oxygen is gas at 25°C and 101.3 kPa.
15. _____ The heat of combustion is the heat released when one mole of substance reacts completely with oxygen gas.
16. _____ The production of acetylene from graphite and hydrogen is an endothermic reaction.
17. _____ Diamond is the standard state for carbon.
18. _____ The standard heat of formation for an element is zero.
19. _____ Formation of NaCl from elemental sodium and chlorine is an exothermic reaction.
20. _____ Standard heats of formation can be used to indirectly calculate the heat of reaction for any reaction that occurs at standard conditions.

**Fill in the Blank:**

21. The standard heat of ________ is the enthalpy change associated with the formation of one mole of a compound from its _______ in their standard states.

**Short Answers:**

22. State Hess’ law.
23. Calculate $\Delta H$ for the reaction: $C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$, from the following data.
24.

   $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l) \quad \quad \quad \Delta H = -1411 \text{ kJ}$

   $C_2H_6(g) + \frac{3}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \quad \quad \quad \Delta H = -1560 \text{ kJ}$

   $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \quad \quad \Delta H = -285.8 \text{ kJ}$
25. Calculate $\Delta H$ for the reaction $4\text{NH}_3(g) + 5\text{O}_2(g) \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g)$, from the following data.

$$N_2(g) + O_2(g) \rightarrow 2\text{NO}(g) \quad \Delta H = -180.5 \text{ kJ}$$
$$N_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g) \quad \Delta H = -91.8 \text{ kJ}$$
$$2\text{H}_2(g) + O_2(g) \rightarrow 2\text{H}_2\text{O}(g) \quad \Delta H = -483.6 \text{ kJ}$$

26. 

27. Calculate $\Delta H^0$ for the reaction $2\text{H}_2(g) + 2\text{C}(s) + \text{O}_2(g) \rightarrow \text{C}_2\text{H}_5\text{OH}(l)$, from the following data.

$$\text{C}_2\text{H}_5\text{OH}(l) + 2\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 2\text{H}_2\text{O}(l) \quad \Delta H = -875 \text{ kJ}$$
$$\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g) \quad \Delta H = -394.51 \text{ kJ}$$
$$\text{H}_2(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{H}_2\text{O}(l) \quad \Delta H = -285.8 \text{ kJ}$$

28. 

29. Calculate $\Delta H^0$ for the reaction $\text{CH}_4(g) + \text{NH}_3(g) \rightarrow \text{HCN}(g) + 3\text{H}_2(g)$, from the following data:

30. 

31. Calculate $\Delta H^0$ for the reaction $2\text{Al}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{AlCl}_3(s)$ from the following data.

32. 

33. Calculate $\Delta H^0$ for the following reactions. Use the data table in lesson 17.4 for thermodynamic information:

1. $N_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g)$
2. $\text{SO}_2(g) + \frac{1}{2}\text{O}_2(g) \rightarrow \text{SO}_3(g)$
3. $\text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s) \quad \Delta H \text{ for } \text{NH}_4\text{Cl} = -314.4 \text{ kJ/mole}$
4. $\text{CaO}(s) + \text{H}_2\text{O}(l) \rightarrow \text{Ca(OH)}_2(s) \quad \Delta H \text{ for } \text{Ca(OH)}_2 = -986.1 \text{ kJ/mole}$
5. $\text{C}_8\text{H}_6(l) + 1\frac{1}{2}\text{O}_2(g) \rightarrow 6\text{C}(s) + 3\text{H}_2\text{O}(l) \quad \Delta H \text{ for } \text{C}_8\text{H}_6 = 49 \text{ kJ/mol}$

---

**Answer Key**

1. a
2. c
3. d
4. b
5. b
6. c
7. d
8. a
9. c
10. b
11. true
24. If two or more thermochemical equations are added together to give a final equation, then the heats of reaction for those equations can also be added together to give a heat of reaction for the final equation.

\[ C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l) \quad \Delta H = -1411. \text{kJ} \]

reverse to get \[ 2CO_2(g) + 3H_2O(l) \rightarrow C_2H_6(g) + 3\frac{1}{2}O_2(g) \quad \Delta H = +1560. \text{kJ} \]

\[ H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H = -285.8 \text{kJ} \]

\[ C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g) \quad \Delta H = -137. \text{kJ} \]

25.

\[ x_2 \quad 2N_2(g) + 2O_2(g) \rightarrow 4NO(g) \quad \Delta H = 2(-180.5 \text{kJ}) \]

reverse \[ x_2 \quad 4NH_3(g) \rightarrow 2N_2(g) + 6H_2(g) \quad \Delta H = -2(-91.8 \text{kJ}) \]

\[ x_3 \quad 6H_2(g) + 3O_2(g) \rightarrow 6H_2O(g) \quad \Delta H = 3(-483.6 \text{kJ}) \]

\[ 4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g) \quad \Delta H = -1628. \text{kJ} \]

26.

reverse \[ x_{\frac{1}{2}} \quad NH_3(g) \rightarrow \frac{1}{2}N_2(g) + \frac{3}{2}H_2(g) \quad \Delta H = -\frac{1}{2}(-91.8 \text{kJ}) \]

reverse \[ \frac{1}{2} \quad CH_4(g) \rightarrow C(s) + 2H_2(g) \quad \Delta H = -1(-74.9 \text{kJ}) \]

\[ \frac{1}{2} \quad H_2(g) + C(s) + \frac{1}{2}N_2(g) \rightarrow HCN(g) \quad \Delta H = \frac{1}{2}(270.3 \text{kJ}) \]

\[ CH_4(g) + NH_3(g) \rightarrow HCN(g) + 3H_2(g) \quad \Delta H = +256.0 \text{kJ} \]

27.

\[ 2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g) \quad \Delta H = -1049. \text{kJ} \]

\[ x_6 \quad 6HCl(g) \rightarrow 6HCl(aq) \quad \Delta H = 6(-74.8 \text{kJ}) \]

\[ x_3 \quad 3H_2(g) + 3Cl_2(g) \rightarrow 6HCl(g) \quad \Delta H = 3(-1845 \text{kJ}) \]

reverse \[ x_2 \quad 2AlCl_3(aq) \rightarrow 2AlCl_3(s) \quad \Delta H = -2(-323. \text{kJ}) \]

\[ 2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s) \quad \Delta H = -6387. \text{kJ} \]
17.4. Hess’s Law

a. \( N_2(g) + 3H_2(g) \rightarrow 2NH_3(g) \)

\[
\Delta H_{\text{rxn}}^\circ = \Delta H_p^\circ - \Delta H_r^\circ \\
= [2NH_3(g)] - [N_2(g) + 3H_2(g)] \\
= [2(-46.1 \text{kJ/mole})] - [0 + 3(0)] \\
= -92.2 \text{kJ/mole} - 0 \text{kJ/mole} \\
= -92.2 \text{kJ/mole}
\]

c. The reaction is exothermic as written with a

d. \( \Delta H_{\text{rxn}}^\circ \)
e. of -92.2 kJ/mole.

f. \( SO_2(g) + \frac{1}{2}O_2(g) \rightarrow SO_3(g) \)

g. \[
\Delta H_{\text{rxn}}^\circ = \Delta H_p^\circ - \Delta H_r^\circ \\
= [SO_3(g)] - [SO_2(g) + \frac{1}{2}O_2(g)] \\
= [(-396.0 \text{kJ/mole})] - [(-296.8 \text{kJ/mole}) + \frac{1}{2}(0)] \\
= -396.0 \text{kJ/mole} + 296.8 \text{kJ/mole} \\
= -99.2 \text{kJ/mole}
\]

h. The reaction is exothermic as written with a

i. \( \Delta H_{\text{rxn}}^\circ \)
j. of -99.2 kJ/mole.

k. \( NH_3(g) + HCl(g) \rightarrow NH_4Cl(s) \)

l. \[
\Delta H_{\text{rxn}}^\circ = \Delta H_p^\circ - \Delta H_r^\circ \\
= [NH_4Cl(s)] - [NH_3(g) + HCl(g)] \\
= [-314.4 \text{kJ/mole}] - [(-46.1 \text{kJ/mole}) + (-92.3 \text{kJ/mole})] \\
= -314.4 \text{kJ/mole} - (-46.1 \text{kJ/mole} + 92.3 \text{kJ/mole}) \\
= -314.4 \text{kJ/mole} - (-138.4 \text{kJ/mole}) \\
= -176.0 \text{kJ/mole}
\]

m. The reaction is exothermic as written with a

n. \( \Delta H_{\text{rxn}}^\circ \)
o. of -176.0 kJ/mole.

p. \( CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s) \)

q. \[
\Delta H_{\text{rxn}}^\circ = \Delta H_p^\circ - \Delta H_r^\circ \\
= [Ca(OH)_2(s)] - [CaO(s) + H_2O(l)] \\
= [-986.1 \text{kJ/mole}] - [(-635.1 \text{kJ/mole}) + (-285.8 \text{kJ/mole})] \\
= -986.1 \text{kJ/mole} - (-920.9 \text{kJ/mole}) \\
= -65.2 \text{kJ/mole}
\]

r. The reaction is exothermic as written with a
s. $\Delta H_{rxn}^\circ$

t. of -65.2 kJ/mole.

u. $C_6H_6(l) + 1\frac{1}{2}O_2(g) \rightarrow 6C(s) + 3H_2O(l)$

v. 

$$\Delta H_{rxn}^\circ = \Delta H_p^\circ - \Delta H_r^\circ$$

$$= [6C(s) + 3H_2O(l)] - [C_6H_6(l) + 1\frac{1}{2}O_2(l)]$$

$$= [6(0) + 3(-285.8 \text{ kJ/mole})] - [(+49.0 \text{ kJ/mole}) + 1\frac{1}{2}(0)]$$

$$= -857.4 \text{ kJ/mole} - (+49.0 \text{ kJ/mole})$$

$$= -906.4 \text{ kJ/mole}$$

w. The reaction is exothermic as written with a

x. $\Delta H_{rxn}^\circ$

y. of -906.4 kJ/mole.
17.5 Thermochemistry

Chapter Test

Name _____________________ Class __________________ Date ________________

Multiple Choice:

1. In a chemical experiment, the surroundings are
   1. the matter in a given space not involved in a reaction
   2. matter in a given space that is involved in the reaction
   3. matter in a given space that is not acted upon
   4. matter in a given space that has already reacted

2. If heat is produced in a chemical reaction, the process is
   1. endothermic
   2. exothermic
   3. endodynamic
   4. exodynamic

3. $\Delta H_f$ is the symbol for
   1. standard heat of combustion
   2. standard heat of state change
   3. standard heat of formation
   4. standard heat of solution

4. A positive enthalpy change means that
   1. heat is absorbed from the surroundings
   2. energy must be put into the reaction
   3. heat is given off to the surroundings
   4. the energy change in the reaction is negligible

5. The heat of fusion for water is
   1. 6.01 kJ/mol
   2. -6.01 kJ/mol
   3. 6.01 J/mol
   4. -6.01 J/mol

6. The heat of fusion deals with
   1. liquid converting to solid
   2. solid converting to gas
   3. solid converting to liquid
   4. liquid converting to gas

7. Calculation of the energy needed to convert 156 water of water at 0°C to solid at 0°C uses the
   1. molar heat of solidification
   2. specific heat of water
3. molar heat of vaporization
4. specific heat of steam

8. The process of converting steam at 125°C to liquid water at 0°C requires ______ steps
   1. five
   2. four
   3. three
   4. two

9. The molar heat of solution for NH₃ is -30.50 kJ/mol. This value tells us that NH₃ will
   1. dissolve slowly in water
   2. generate a great deal of heat when dissolving
   3. increase the specific heat of the solution
   4. decrease solvent temperature while dissolving

10. The standard state for chlorine is
    1. gas at 25°C and 101.3 kPa
    2. gas at 20°C and 101.3 kPa
    3. liquid at 25°C and 101.3 kPa
    4. liquid at 20°C and 101.3 kPa

True/False:

11. _____ 1000 calories = 1 Calorie
12. _____ The sign of q is positive for an exothermic process
13. _____ Water has a low specific heat.
14. _____ Chemical reactions in the lab are run with controlled pressure.
15. _____ A lid is used to prevent outside contamination in a reaction system.
16. _____ The units for ∆H° are kJ/mol
17. _____ The conversion of vapor directly to solid is known as sublimation.
18. _____ The standard heat of formation of an element depends on the element.
19. _____ Energy is released when a liquid is converted to a solid.
20. _____ Hₛ is the symbol for specific heat.

Short Answers:

21. How many calories are needed to raise 125 g water 10.7°C?
22. A chemical reactions released 895 calories of energy. How many joules were released?
23. Sunlight shining on a sheet of glass with a mass of 115 g showed a 5.2°C increase in temperature when 502.3 J of heat is supplied. What is the specific heat of the glass?
24. In an experiment, 25.0 mL of 0.50 M HBr at 20.0°C is added to 25.0 mL of 0.50 M KOH at 20.0°C in a foam cup calorimeter. As the reaction occurs, the temperature of the solution rises to 32.0°C. Calculate the enthalpy change (∆H) in kJ for this reaction. Assume the densities of the solutions are 1.00 g/mL and that their specific heats are the same as that of pure water.
25. H₂(g) + S(s) + 2O₂(g) → H₂SO₄(l) \( \Delta H^\circ = -811 \text{ kJ mol}^{-1} \)
26. What is the energy released if 47.3 g sulfur are reacted to form sulfuric acid?
27. Calculate the amount of energy absorbed when 89.7 grams of liquid ethanol at 78.4°C is converted to the vapor state at the same temperature.
28. How much energy is required to convert 118 g water at 55°C to steam at 100°C?
29. How many joules of energy are absorbed when 76.2 g KI (heat of solution = 20.33 kJ/mol) are dissolved in 750 mL water? If the initial water temperature is 22.7°C, what will the final temperature be?
30. Use the thermochemical equations shown below to determine the enthalpy for the reaction \( \text{CH₃COOH}(l) \rightarrow 2\text{C}(s) + 2\text{H}_2(g) + \text{O}_2(g) \)
31. 

\[
2\text{CO}_2(g) + 2\text{H}_2\text{O}(l) \rightarrow \text{CH}_3\text{COOH}(l) + 2\text{O}_2(g) \quad \Delta H^\circ = 3484 \text{ kJ}
\]

\[amp;C(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g) \quad \Delta H^\circ = -1576 \text{ kJ}\]

\[2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) \quad \Delta H^\circ = -2288 \text{ kJ}\]

32. When potassium chloride reacts with oxygen under the right conditions, potassium chlorate is formed:

33. \[2\text{KCl} + 3\text{O}_2 \rightarrow 2\text{KClO}_3\]

34. Given that the heat of formation of potassium chloride is -36 kJ/mol and the heat of formation of potassium chlorate is -391 kJ/mol, determine the heat of reaction.

---

**Answer Key**

1. a
2. b
3. c
4. a
5. a
6. c
7. a
8. c
9. b
10. b
11. true
12. false
13. false
14. true
15. false
16. true
17. false
18. false
19. true
20. false

21. 

\[\text{calories} = 125 \text{ g} \times 10.7^\circ C \]

\[\text{calories} = \frac{100 \text{ g} \times 10.7^\circ C \times \text{cal}}{g \cdot ^\circ C} = 1007 \text{ calories}\]

22. 895 calories $\times 4.184 \text{ J/cal} = 3745 \text{ J}$

23. \[q = m \times c_p \times \Delta T \quad \text{so} \quad c_p = \frac{q}{m \times \Delta T} = \frac{502.3 \text{ J}}{115 \text{ g} \times 5.2^\circ C} = 0.84 \text{ J/g} \cdot ^\circ C\]

24. \[q = m \times c_p \times \Delta T = 50 \text{ g} \times 4.184 \text{ J/g} \cdot ^\circ C \times 12^\circ C = 2510.4 \text{ J}\]

25. \[47.3 \text{ g} \div 32.1 \text{ g/mol} = 1.47 \text{ moles } S\]

26. \[1.47 \text{ moles} \times (-811 \text{ kJ/mol}) = 1192.2 \text{ kJ}\]

27. ethanol molar mass = 46 g/mol

28. \[89.7 \text{ g} \div 46 \text{ g/mol} = 1.95 \text{ moles}\]

29. \[1.95 \text{ moles} \times 43.5 \text{ kJ/mol} = 84.8 \text{ kJ}\]

a. 118 g water at 55°C to 118 g water at 100°C \(\Delta T = 45^\circ C\)
b. \( q = m \times c_p \times \Delta T = 118 \text{ g} \times 4.184 \text{ J/g} \cdot ^\circ \text{C} \times 45^\circ \text{C} = 22217 \text{ J} = 22.22 \text{ kJ} \)

c. 118 g water at 100°C to steam at 100°C

d. \( 118 \text{ g} \div 18.01 \text{ g/mole} = 6.55 \text{ moles} \)

e. \( 6.55 \text{ moles} \times 40.7 \text{ kJ/mol} = 266.6 \text{ kJ} \)

f. total energy = 22.22 kJ + 266.6 kJ = 288.8 kJ

28. molar mass of KI = 166 g/mole

29. 76.2 g \div 166 \text{ g/mole} = 0.46 \text{ moles} \)

30. 20.33 kJ/mole \times 0.46 \text{ moles} = 9.35 \text{ kJ} = 9350 \text{ J} \)

31. \( q = m \times c_p \times \Delta T \)

32. \( \Delta T = \frac{q}{m \times c_p} = \frac{9350 \text{ J}}{166 \text{ g} \times 4.18 \text{ J/g} \cdot ^\circ \text{C}} = 13.5^\circ \text{C} \)

33. \( 22.7^\circ \text{C} - 13.5^\circ \text{C} = 9.2^\circ \text{C} \)

34. final temperature

35.

reverse \( \text{CH}_3\text{COOH}(l) + 2\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 2\text{H}_2\text{O}(l) \) \( \Delta H^\circ = -3484 \text{ kJ} \)

reverse and double \( 2\text{CO}_2(g) \rightarrow 2\text{C}(s) + 2\text{O}_2(g) \) \( \Delta H^\circ = +3152 \text{ kJ} \)

reverse \( 2\text{H}_2\text{O}(l) \rightarrow 2\text{H}_2(g) + \text{O}_2(g) \) \( \Delta H^\circ = +2288 \text{ kJ} \)

\( \text{CH}_3\text{COOH}(l) \rightarrow 2\text{C}(s) + 2\text{H}_2(g) + \text{O}_2(g) \) \( \Delta H^\circ = 1956 \text{ kJ} \)

36.

\( \Delta H_{rxn} = \Delta H_f(products) - \Delta H_f(reactants) \)

\( \Delta H_{rxn} = [2(-391 \text{ kJ/mol})] - [2(-436 \text{ kJ/mol})] \)

\( \Delta H_{rxn} = -782 \text{ kJ/mol} + 872 \text{ kJ/mol} \)

\( \Delta H_{rxn} = 90. \text{ kJ/mol} \)
## Chapter Outline

<table>
<thead>
<tr>
<th>Section</th>
<th>Title</th>
</tr>
</thead>
<tbody>
<tr>
<td>18.1</td>
<td>Rates of Reactions</td>
</tr>
<tr>
<td>18.2</td>
<td>Rate Laws</td>
</tr>
<tr>
<td>18.3</td>
<td>Reaction Mechanisms</td>
</tr>
<tr>
<td>18.4</td>
<td>Kinetics</td>
</tr>
</tbody>
</table>
Lesson Quiz

Multiple Choice:

1. Rate is defined as
   1. molarity/minute
   2. moles/second
   3. molals/minute
   4. molarity/second

2. Rate is another word for
   1. time
   2. change
   3. speed
   4. distance

3. One of the following is not a condition for a product to form
   1. collision at correct orientation
   2. proper size of molecules
   3. sufficient kinetic energy
   4. energetic collision

4. No rearrangement of atoms occurs during an
   1. ineffective collision
   2. excessive collision
   3. asymmetric collision
   4. effective collision

5. In a potential energy diagram for an exothermic reaction
   1. the potential energy of products is greater than that of reactants
   2. the potential energy of reactants and products is equal
   3. the potential energy of reactants is greater than that of products
   4. the potential energy valley is at a maximum

6. Activation energy
   1. is shown as a peak on the potential energy graph
   2. is the energy released by an exothermic reaction
   3. is the energy stored by an endothermic reaction
   4. is the difference between energy of reactants and energy of products

7. To include a catalyst in a reaction, we must
   1. write a separate procedure
   2. write a separate equation
3. include it in the chemical equation
4. write a footnote to the procedure

8. An increase in pressure will increase the rate of a reaction in the _______ phase
   1. solid
   2. gas
   3. liquid
   4. solution

9. An increase in collision rate in a solution can be accomplished by
   1. an increase in particle size
   2. an increase in pressure
   3. an increase in temperature
   4. an increase in solvent volume

10. The role of a catalyst is to
    1. lower activation energy
    2. increase activation energy
    3. lower \( \Delta H \) of products
    4. increase \( \Delta H \) of products

True/False:

11. _____ A catalyst is changed by the reaction it catalyzes.
12. _____ An increase in temperature raises the collision frequency of molecules in a reaction.
13. _____ [A] refers to the number of molals of A.
14. _____ In a rate equation, \( \Delta \) indicates a change in.
15. _____ All reactions occur at higher temperatures.
16. _____ Lower vibrational energies produce more bond breaking.
17. _____ Little is known about the structures of most activated complexes.
18. _____ More energetic molecules have more forceful collisions.
19. _____ An increase in gas pressure does not affect the concentration of the gas.
20. _____ A potential energy diagram is sometimes called a reaction progress curve.

Fill in the Blank:

21. A reaction rate is the change in _________ of a reactant or product with _____.
22. Collision ______ is a set of principles based around the idea that _____ particles form _____ when they collide with one another, but only when those collisions have enough ________ energy and the correct ______________ to cause a reaction.
23. The __________ energy for a reaction is the _________ energy that colliding particles must have in order to ________ a reaction.
24. A _______ _______ diagram shows the change in the potential energy of a _____ as _____ are converted into ______.
25. The total potential energy of the system __________ for the endothermic reaction as the system ________- energy from the __________.
26. An ________ complex is an ________ arrangement of atoms that exists ______ at the ________ of the activation energy ________.
27. When ______ particles are present in a given amount of ______, a ________ number of ________ will naturally occur between those particles.
28. ______ gas ________ leads to a greater ______ of ________ between reacting particles.
29. An ________ in the ______ area of a reactant ________ the rate of a reaction.
30. A _______ is a substance that ____________ the _______ of a chemical reaction without being used up in the reaction.

31. Identify each part of the following diagram:

32.

33. Identify the profile for an exothermic reaction.

34. Profile A  Profile B

Answer Key

1. d
2. c
3. b
4. a
5. c
6. a
7. c
8. b
9. c
10. a
11. false
18.1. Rates of Reactions

12. true
13. false
14. true
15. false
16. false
17. true
18. true
19. false
20. true
21. concentration, time
22. theory, reactant, products, kinetic, orientation
23. activation, minimum, undergo
24. potential energy, system, reactants
25. increases, absorbs, surroundings
26. activated, unstable, momentarily, peak, barrier
27. more, space, greater, collisions
28. higher, pressure, frequency, collisions
29. increase, surface, increases
30. catalyst, increases, rate
   1. potential energy of reactants
   2. activation energy
   3. activated complex
   4. product potential energy
31. Profile B represents the exothermic reaction.
18.2 Rate Laws

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. The rate law for a given reaction is \( \text{rate} = k[A][B]^2 \). The reaction is _____ order overall.
   1. zero
   2. first
   3. second
   4. third

2. The reaction in question one is ______ order with regard to A.
   1. zero
   2. first
   3. second
   4. third

3. The reaction in question one is _____ order with regard to B.
   1. zero
   2. first
   3. second
   4. third

4. In the reaction \( A + B \rightarrow C \), the rate law is \( \text{rate} = k[A] \). Therefore, the reaction is _____ order with regard to B.
   1. zero
   2. first
   3. second
   4. third

5. A large value for \( k \) means the reaction is
   1. relatively fast
   2. biphasic
   3. relatively slow
   4. incomplete

6. In the reaction \( A + B \rightarrow C + D \), when the concentration of A is doubled at constant [B], the rate doubles. This data indicates that the reaction is ______ order for A.
   1. zero
   2. first
   3. second
   4. third

7. For the reaction in problem 6, when [B] is doubled at constant [A], the reaction rate is four times what it was. This data indicates that the reaction is ______ order for B.
   1. zero
   2. first
   3. second
   4. third
18.2. Rate Laws

1. zero
2. first
3. second
4. third

8. For the reaction in problem 6, when \([B]\) is doubled at constant \([A]\), the reaction rate is unchanged. This data indicates that the reaction is ______ order for \(B\).
   1. zero
   2. first
   3. second
   4. third

9. In the hypothetical reaction \(A + B \rightarrow C + D\), if we double the amount of \(A\),
   1. the rate will be unchanged
   2. the rate will decrease by half
   3. the rate will be doubled
   4. we cannot predict the rate with the available data

10. In a given reaction, when the concentration of \(B\) is doubled, the rate is increased four-fold. We can write the statement
    1. rate \(\propto [B]\)
    2. rate \(\propto [B]^2\)
    3. rate \(\propto [B]^4\)
    4. rate \(\propto [B]^0\)

True/False:

11. _____ The rate of a reaction decreases as the concentration of reactants decreases.
12. _____ The rate of a reaction is not affected by collision frequency.
13. _____ The rate law for an equation can be determined by visual inspection.
14. _____ The value for a rate constant depends on the temperature.
15. _____ Real reactions do not always proceed by single-step mechanisms.
16. _____ In the reaction \(A + 2B \rightarrow C\), if we double \([B]\), we quadruple the rate of the reaction.
17. _____ The units for the specific rate constant are \(\text{M/sec}\).
18. _____ The specific rate constant must be determined experimentally.
19. _____ The value of the rate constant tells us how fast the reaction will proceed.
20. _____ In a first-order reaction, there is only one reactant.

Fill in the Blank:

21. A rate ____ is an expression showing the relationship of the ______ rate to the ______ of each reactant.
22. A ______-order reaction is a reaction in which the rate is ______ proportional to the concentration of a ______ reactant.
23. Reaction kinetics are based on initial rates of reaction, measured as soon as possible after the reaction is initiated. Why can we not use rate data after the reaction has run for several minutes?

24. For each of the following sets of data, determine the rate law expression, the reaction order, and the specific rate constant:
25. a.
26. \(A_2 + B_2 \rightarrow 2AB\)
### TABLE 18.1:

<table>
<thead>
<tr>
<th>Exp #</th>
<th>( [A_2] )</th>
<th>( [B_2] )</th>
<th>Rate (M/sec)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.001</td>
<td>0.001</td>
<td>0.01</td>
</tr>
<tr>
<td>2</td>
<td>0.001</td>
<td>0.002</td>
<td>0.02</td>
</tr>
<tr>
<td>3</td>
<td>0.001</td>
<td>0.003</td>
<td>0.03</td>
</tr>
<tr>
<td>4</td>
<td>0.001</td>
<td>0.004</td>
<td>0.04</td>
</tr>
<tr>
<td>5</td>
<td>0.002</td>
<td>0.004</td>
<td>0.16</td>
</tr>
<tr>
<td>6</td>
<td>0.003</td>
<td>0.004</td>
<td>0.36</td>
</tr>
</tbody>
</table>

1. b.
2. \( F + G \rightarrow H \)

### TABLE 18.2:

<table>
<thead>
<tr>
<th>Exp #</th>
<th>( [F] )</th>
<th>( [G] )</th>
<th>Rate ((M/sec))</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.01</td>
<td>0.4</td>
<td>0.02</td>
</tr>
<tr>
<td>2</td>
<td>0.02</td>
<td>0.4</td>
<td>0.04</td>
</tr>
<tr>
<td>3</td>
<td>0.03</td>
<td>0.4</td>
<td>0.06</td>
</tr>
<tr>
<td>4</td>
<td>0.1</td>
<td>0.2</td>
<td>5</td>
</tr>
<tr>
<td>5</td>
<td>0.1</td>
<td>0.4</td>
<td>10</td>
</tr>
<tr>
<td>6</td>
<td>0.1</td>
<td>0.6</td>
<td>15</td>
</tr>
</tbody>
</table>

1. c.
2. \( F + G \rightarrow H \)

### TABLE 18.3:

<table>
<thead>
<tr>
<th>Exp #</th>
<th>( [F] )</th>
<th>( [G] )</th>
<th>Rate ((M/sec))</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.01</td>
<td>0.4</td>
<td>0.02</td>
</tr>
<tr>
<td>2</td>
<td>0.02</td>
<td>0.4</td>
<td>0.16</td>
</tr>
<tr>
<td>3</td>
<td>0.03</td>
<td>0.4</td>
<td>0.54</td>
</tr>
<tr>
<td>4</td>
<td>0.1</td>
<td>0.2</td>
<td>5</td>
</tr>
<tr>
<td>5</td>
<td>0.1</td>
<td>0.4</td>
<td>20</td>
</tr>
<tr>
<td>6</td>
<td>0.1</td>
<td>0.6</td>
<td>45</td>
</tr>
</tbody>
</table>

25. For the reaction \( A + B \rightarrow C \), the rate law is \( rate = k[A]^2[B] \). Fill in the blanks of the following table to show data that would give the indicated rate law.

### TABLE 18.4:

<table>
<thead>
<tr>
<th>Experiment</th>
<th>( [A] )</th>
<th>( [B] )</th>
<th>Rate ((M/sec))</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.1</td>
<td>0.1</td>
<td>1</td>
</tr>
<tr>
<td>2</td>
<td>0.1</td>
<td>0.2</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>0.1</td>
<td>0.3</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>0.2</td>
<td>0.1</td>
<td></td>
</tr>
<tr>
<td>5</td>
<td>0.4</td>
<td>0.1</td>
<td></td>
</tr>
</tbody>
</table>
Answer Key

1. d
2. b
3. c
4. a
5. a
6. b
7. c
8. a
9. d
10. b
11. true
12. false
13. false
14. true
15. true
16. false
17. true
18. true
19. true
20. true
21. law, reaction, concentrations
22. first, directly, single
23. After several minutes, the concentrations of reactants have often changed significantly. We need accurate measurements of concentrations, which can only be assumed at the beginning of the reaction.

**Table 18.5:**

<table>
<thead>
<tr>
<th>Experiment</th>
<th>[A]</th>
<th>[B]</th>
<th>Rate (M/sec)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.1</td>
<td>0.1</td>
<td>1</td>
</tr>
<tr>
<td>2</td>
<td>0.1</td>
<td>0.2</td>
<td>2</td>
</tr>
<tr>
<td>3</td>
<td>0.1</td>
<td>0.3</td>
<td>3</td>
</tr>
<tr>
<td>4</td>
<td>0.2</td>
<td>0.1</td>
<td>4</td>
</tr>
<tr>
<td>5</td>
<td>0.4</td>
<td>0.1</td>
<td>5</td>
</tr>
</tbody>
</table>
Lesson Quiz

Name___________________ Class________________ Date ________

Multiple Choice:

1. The overall reaction tells us
   1. the individual steps in the process
   2. the materials needed for the reaction
   3. how the reaction proceeds
   4. the rate equation for the reaction

2. In most reactions, a _____ ______ is formed
   1. reaction complex
   2. intermediate complex
   3. reaction intermediate
   4. intermediary complex

3. A _______ reaction involves three reacting molecules in one elementary step
   1. unimolecular
   2. trimolecular
   3. bimolecular
   4. termolecular

4. In the overall reaction \(2NO + O_2 \rightarrow 2NO_2\), a reaction that shows an intermediate is called
   1. intermediary reaction
   2. intermediate reaction
   3. elementary reaction
   4. elementary process

5. In a unimolecular reaction, product formation
   1. increases linearly with \([A]\)
   2. increases non-linearly with \([A]\)
   3. is independent of \([A]\)
   4. increases linearly with \(\frac{1}{[A]}\).

6. In a potential energy diagram, the well between two peaks represents the
   1. activated complex
   2. reaction intermediate
   3. \(\Delta H\) for the system
   4. \(\Delta H\) for the elementary step

True/False:

7. _____ The overall reaction does not include catalysts
8. _____ Reaction intermediates help determine the order of the reaction.
9. _____ A unimolecular reaction can only be first order.
10. _____ The overall enthalpy change in a potential energy diagram is not influenced by the activated complex.
11. _____ A catalyst can be consumed in one step in a reaction mechanism.
12. _____ Activated complexes can be detected and characterized.

Fill in the Blank:

13. An _____ step is one step in a ________ of simple reactions that show the _____ of a _______ at the _______ level.
14. A reaction ________ is the _______ of _______ steps that together comprise an entire chemical ________.
15. An ________ is a ________ which appears in the _________ of a reaction, but not in the overall ______- ______ equation.
16. The _______ of an _______ step is the total number of _______ molecules in that step.
17. The reaction ________ for each _________ in an elementary step is equal to its _________ coefficient in the equation for that step.
18. The rate-_______ step is the _______ step in the reaction mechanism.
19. For the reaction $NO_2 + CO \rightarrow NO + CO_2$ (all in gas phase), the experimental rate $= k[NO_2]^2$. Propose a mechanism and identify the rate-determining step.
20. Propose a mechanism for the reaction $2NO_2 + F_2 \rightarrow 2NO_2F$ (all in gas phase), with an experimental rate law of $k[NO_2][F_2]$ and identify the rate-limiting step.
21. In the reaction $O_3 + O \rightarrow 2O_2$, a two-step mechanism is proposed:
   22. Step one:
      23. $O_3 + NO \rightarrow NO_2 + O_2$
   24. (slow) Step two:
      25. $NO_2 + O \rightarrow NO + O_2$
   26. (fast) Identify the catalyst and the intermediate. Explain your answer.
27. The composition of $H_2O_2$ in aqueous solution to form water plus oxygen is catalyzed by $Br_2$. The following two-step mechanism has been proposed:
   28. Step One:
      29. $H_2O_2 + Br_2 \rightarrow 2Br^- + 2H^+ + O_2$
   30. Step Two:
      31. $H_2O_2 + 2Br^- + 2H^+ \rightarrow Br_2 + 2H_2O$
   32. Write the overall reaction and explain why the mechanism is consistent with it.

Answer Key

1. b
2. c
3. d
4. b
5. a
6. b
7. false
8. true
9. true
10. false
11. true
12. false
13. elementary, series, progress, reaction, molecular
14. mechanism, sequence, elementary, reaction
15. intermediate, species, mechanism, balanced
16. molecularity, elementary, reactant
17. order, reactant, stoichiometric
18. determining, slowest
19. Step one: \( NO_2 + NO_2 \rightarrow NO_3 + NO \)
20. Step two:
21. \( NO_3 + CO \rightarrow NO_2 + CO_2 \)
22. Rate-determining step is step one since its rate equation fits the experimental one.
23. Step one: \( NO_2 + F_2 \rightarrow NO_2F + F \)
24. Step two:
25. \( NO_2 + F \rightarrow NO_2F \)
26. Rate-determining step is step one – fits the experimental rate equation
27. \( NO \) is the catalyst since it appears at both beginning and end of process.
28. \( NO_2 \)
29. is an intermediate because it does not appear at the beginning or end of the process.
30. Overall reaction: \( 2H_2O_2 \xrightarrow{Br_2} 2H_2O + O_2 \)
31. \( Br_2 \)
32. is used in step one and regenerated in step two, playing the role of the catalyst.
18.4. Kinetics

Chapter Test

Name___________________ Class______________ Date________

Multiple Choice:

1. Rate is defined as
   1. \( \Delta A \div \Delta T \)
   2. \( \Delta A \times \Delta T \)
   3. \( \Delta T \div \Delta A \)
   4. \( \Delta A - \Delta T \)

2. Chemical reactions are measured in
   1. M/min
   2. m/min
   3. M/sec
   4. m/sec

3. Ineffective collisions produce
   1. smaller molecules
   2. fewer molecules
   3. no reaction
   4. slower reaction

4. All of the following will produce an increase in reactant collision rate except
   1. an increase in concentration
   2. an increase in solvent
   3. an increase in temperature
   4. a decrease in particle size

5. Addition of a catalyst to a reaction
   1. increases the activation energy
   2. increases the \( \Delta H \) for the reaction
   3. decreases the activation energy
   4. decreases the \( \Delta H \) for the reaction

6. The specific rate constant
   1. depends on the type of reaction
   2. can be calculated from the stoichiometry of the equation
   3. must be determined experimentally
   4. is independent of the rate law

7. Rate constants are determined by using
   1. initial rates
   2. final rates
3. multiple rates
4. incremental rates

8. In the reaction $A + B \rightarrow C$, when the concentration of $B$ is doubled, the rate is not affected. The reaction is _______ - order for $B$.
   1. zero
   2. first
   3. second
   4. third

9. A reaction intermediate is
   1. shown as a reactant in the equation
   2. shown as a product in the equation
   3. written over the arrow in the equation
   4. not included in the equation

10. A potential energy diagram shows all of the following except
    1. reactant potential energy
    2. activation energy
    3. reaction order
    4. product potential energy

True/False:

11. _____ $[A] = \text{molar concentration of } A$.
12. _____ Temperature does not affect rate constant values.
13. _____ If a reaction is second-order for reactant $B$, the rate will triple when $[B]$ is doubled.
14. _____ Catalysts are included in the equation for an overall reaction.
15. _____ The identity of reaction intermediates does not influence determination of reaction order.
16. _____ Lower vibrational energies produce more bond breakage.
17. _____ Activated complexes are easy to characterize.
18. _____ A reaction complex is the same as an activated complex.
19. _____ The $\Delta H$ of the activated complex affects the overall $\Delta H$ of the reaction.
20. _____ The energy trough in a potential energy diagram represents an intermediate in the reaction.

Fill in the Blank:

21. Define the following terms:
   1. reaction rate
   2. activation energy
   3. catalyst
   4. specific rate constant
   5. reaction mechanism

22. Calculate the reaction rates for the following reactions in M/sec:
   1. initial concentration: 4.23 M concentration after 3.5 minutes: 2.7M
   2. initial concentration: 2.14 M concentration after 0.7 minutes: 1.37M

23. The reaction between hydrogen gas and oxygen gas forms water. Which of the following conditions will lead to the lowest rate of reaction? The highest?
   1. gas pressures 0.01 atm each, temperature -12°C
   2. gas pressures 0.01 atm each, temperature 45°C
   3. gas pressures 1.5 atm each, temperature -12°C
4. gas pressures 1.5 atm each, temperature 45°C

24. The following experimental data were obtained for the reaction at 250 K,
25. \( F_2 + 2ClO_2 \rightarrow 2FClO_2 \)

<table>
<thead>
<tr>
<th>Table 18.6:</th>
</tr>
</thead>
<tbody>
<tr>
<td>[ F_2 ] M</td>
</tr>
<tr>
<td>----------------</td>
</tr>
<tr>
<td>0.10</td>
</tr>
<tr>
<td>0.10</td>
</tr>
<tr>
<td>0.20</td>
</tr>
</tbody>
</table>

1. Write a rate law consistent with this data.

25. The following data were collected for the reaction
26. \( P_4 + 6H_2 \rightarrow 4PH_3 \)

<table>
<thead>
<tr>
<th>Table 18.7:</th>
</tr>
</thead>
<tbody>
<tr>
<td>[ P_4 ] M</td>
</tr>
<tr>
<td>----------------</td>
</tr>
<tr>
<td>0.0110</td>
</tr>
<tr>
<td>0.0110</td>
</tr>
<tr>
<td>0.0220</td>
</tr>
</tbody>
</table>

1. What is the reaction order for
2. \( P_4 \)
3. \( H_2 \)
4. \( H_2 \)
5. Write the rate law.

26. Given the reaction:
27. \( 4HBr + O_2 \rightarrow 2H_2O + 2Br_2 \)
28. a. Would you expect this reaction to take place in a single step? ______ Why or why not? _____________

29. b. This reaction is thought to take place by means of the following mechanism: Step 1: 
30. \( HBr + O_2 \rightarrow HOOBr \)
31. (slow) Step 2: 
32. \( HBr + HOOBr \rightarrow 2HOBr \)
33. (fast) Step 3: 
34. \( 2HBr + 2HOBr \rightarrow 2H_2O + 2Br_2 \)
35. (fast) If this mechanism is valid, what would the rate law for the reaction look like?
36. c. Identify two reaction intermediates in this mechanism.

Answer Key

1. a
2. c
3. c
1. The change in concentration of a reactant or product with time.
2. The minimum energy that colliding particles must have in order to undergo a reaction.
3. A substance that increases the rate of a chemical reaction without being used up in the reaction
4. The proportionality constant relating the rate of the reaction to the concentrations of the reactants.
5. The sequence of elementary steps that together comprise an entire chemical reaction.

\[
1. \quad \text{rate} = \frac{\Delta A}{\Delta t} = \frac{1.53 \text{ M}}{3.5 \text{ min}} \times \frac{1 \text{ min}}{60 \text{ sec}} = 0.0072 \text{ M/sec}
\]
\[
2. \quad \text{rate} = \frac{\Delta A}{\Delta t} = \frac{0.77 \text{ M}}{0.7 \text{ min}} \times \frac{1 \text{ min}}{60 \text{ sec}} = 0.018 \text{ M/sec}
\]

21. The lowest rate will be at the lowest pressure plus lowest temperature or selection a.
22. The highest rate will be at the highest pressure plus the highest temperature or election d.
23. A doubling of \([F_2]\) gives a four-fold increase in rate, so the reaction is second order for \([F_2]\). A four-fold increase in \(ClO_2\) gives a four-fold increase in rate, so the reaction is first-order for \(ClO_2\). The rate law is \(\text{rate} = k[F_2]^2[ClO_2]\).
24. The rate is unaffected by \([P_4]\), reaction is zero order for \(P_4\). When \([H_2]\) is doubled, the reaction rate is doubled, reaction is first-order for \(H_2\). Rate law: \(\text{rate} = k[H_2]\).

1. The reaction as written has five molecules involved, a very unlikely situation of all five need to collide simultaneously.
2. A possible rate law would be \(\text{rate} = k[HBr][O_2]\)
3. Reaction intermediates are HOBr and HOOBr.
CHAPTER 19
Equilibrium Assessments

Chapter Outline

19.1 The Nature of Equilibrium
19.2 Le Châtelier’s Principle
19.3 Solubility Equilibrium
19.4 Equilibrium
Lesson Quiz

Name___________________ Class______________ Date________

*Multiple Choice:*

1. The symbol $\rightleftharpoons$ indicates
   1. reversible reaction
   2. formation of a precipitate
   3. biphasic reaction
   4. unbalanced equation

2. In a reversible reaction, all of the following are true except
   1. the reactant concentration decrease with time
   2. the product concentration increases with time
   3. the concentrations of reactants and products become equal
   4. the forward rate equals the reverse rate

3. An equilibrium expression includes only
   1. concentrations of solids
   2. concentrations of gases and aqueous solutions
   3. concentrations of liquid water
   4. concentrations of pure liquids

4. The units for $K_{eq}$ are
   1. moles/L
   2. grams/L
   3. none
   4. depends on the equation

5. A value of $K_{eq}$ greater than 1 means
   1. products favored over reactants
   2. reactants favored over products
   3. an equilibrium situation
   4. reaction has not yet come to completion

6. For the reaction $aA + bB \rightleftharpoons cC + dD$ the equilibrium expression is written
   1. $a[A]b[B] = c[C]d[D]$
   2. $\frac{[C][D]}{[A][B]}$
   3. $\frac{a[A]b[B]}{c[C]d[D]}$
   4. $\frac{[C][D]}{[A][B]}$

7. Ammonium carbonate decomposes to form
   1. ammonium carbonate
2. ammonia + carbon dioxide
3. ammonia + carbon monoxide
4. ammonium nitrate + carbon dioxide

True/False:
8. _____ Not all equilibrium reactions are reversible.
9. _____ The equilibrium position depends which materials are added first.
10. _____ Equilibrium reactions must be run in a closed system.
11. _____ Chemical equilibrium is a dynamic process.
12. _____ The value for $K_{eq}$ of a system depends on which direction the reaction is written.
13. _____ A phase equilibrium exists when a substance is in equilibrium between two states.
14. _____ A change in temperature changes the rate of forward and reverse reactions.

Fill in the blanks:
15. A _________ reaction is a reaction in which the ________ of reactants to products and the conversion of products to reactants occur __________.
16. Chemical _________ is the state of a system in which the ______ of the forward reaction is ______ to the rate of the reverse reaction.
17. The ________ position is a property of the particular ________ reaction and does not depend upon the ________ concentrations of the reactants and products.
18. The equilibrium ________ (K$_{eq}$) is the ratio of the product of the concentrations of the products to the mathematical product of the concentrations of the reactants for a ________ that is at equilibrium.
19. Since the product concentrations are in the ______ of the equilibrium expression, a $K_{eq} > 1$ means that the products are ______ over the reactants.
20. Write equilibrium constant expressions for the following reactions:
   1. $2SO_2 + O_2 \rightleftharpoons 2SO_3$
   2. $C_2H_6 \rightleftharpoons C_2H_4 + H_2$
   3. $As_4O_6(S) + 6C(S) \rightleftharpoons As_4(g) + 6CO(g)$
   4. $SnO_2(S) + 2CO(g) \rightleftharpoons Sn(S) + 2CO_2(g)$
21. Calculate the $K_{eq}$ of the following reaction
22. $S_2(g) + 2H_2(g) \rightleftharpoons 2H_2S(g)$
23. 
   \[ [H_2] = 2.16 \, M \]
   \[ S_2 \]
   \[ = 0.3 \, M \]
   \[ H_2S \]
   \[ = 0.5 \, M \]
24. Six moles of $SO_2(g)$ and four moles of $O_2(g)$ are introduced into a 1.00 L reaction vessel and allowed to react to form $SO_3(g)$. At equilibrium, the vessel contains four moles of $SO_3(g)$. Calculate $K_{eq}$ for this reaction.
25. The equilibrium constant for the reaction below is 0.11. Calculate all equilibrium concentrations if 0.33 mol of iodine chloride gas is placed in a 1.00 L vessel and allowed to come to equilibrium.
26. $2ICl(g) \rightleftharpoons I_2(g) + Cl_2(g)$
27. The following reaction has a $K_{eq}$ value of 85.0 at 460°C:
28. $SO_2(g) + NO_2(g) \rightleftharpoons NO(g) + SO_3(g)$
29. If a mixture of sulfur dioxide and nitrogen dioxide is prepared, each with an initial concentration of 0.100 mol/L, calculate the equilibrium concentrations of nitrogen dioxide and nitrogen monoxide at this temperature.
Answer Key

1. a
2. c
3. b
4. a
5. a
6. d
7. b
8. false
9. false
10. true
11. true
12. true
13. true
14. true
15. reversible, conversion, simultaneously
16. equilibrium, rate, equal
17. equilibrium, reversible, initial
18. constant, mathematical, reaction
19. product, numerator, favored

\[
1. \quad K_{eq} = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]}
\]

\[
2. \quad K_{eq} = \frac{[\text{C}_2\text{H}_6]^2}{[\text{C}_2\text{H}_4][\text{H}_2]}
\]

\[
3. \quad K_{eq} = [\text{As}_4][\text{CO}]^6
\]

\[
4. \quad K_{eq} = \frac{[\text{CO}_2]^2}{[\text{CO}]^2}
\]

\[
20. \quad K_{eq} = \frac{[\text{H}_2\text{S}]^2}{[\text{S}_2][\text{H}_2]^2} = \frac{(0.5 \text{ M})^2}{(0.3 \text{ M})^2(2.16 \text{ M})^2} = 0.178
\]

21. \[2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)\]

Table 19.1:

<table>
<thead>
<tr>
<th>I</th>
<th>6.00</th>
<th>4.00</th>
<th>0</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>$-2x$</td>
<td>$-x$</td>
<td>$+2x$</td>
</tr>
<tr>
<td>E</td>
<td>$6.00 - 2x$</td>
<td>$4.00 - x$</td>
<td>$2x = 4.00$</td>
</tr>
</tbody>
</table>

At equilibrium \([\text{SO}_3] = 4.00 \text{ M}\)

Therefore \(x = 2.00 \text{ M}\)

\([\text{SO}_2] = 2.00 \text{ M}\)

\(O_2 \quad = 2.00 \text{ M}\)

\[
K_{eq} = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]} = \frac{(4.00)^2}{(2.00)^2(2.00)} = 2.00
\]

23. \[2\text{ICl}(g) \rightleftharpoons \text{I}_2(g) + \text{Cl}_2(g)\] \(K_{eq} = 0.11\)
24. $2ICl(g) \rightleftharpoons I_2(g) + Cl_2(g)$

**Table 19.2:**

<table>
<thead>
<tr>
<th>I</th>
<th>0.33</th>
<th>0</th>
<th>0</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>$-2x$</td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
<td>E</td>
<td>$0.33 - 2x$</td>
<td>$x$</td>
<td>$x$</td>
</tr>
</tbody>
</table>

$K_{eq} = \frac{[I_2][Cl_2]}{[ICl]^2}$ or $0.11 = \frac{(x)(x)}{(0.33 - 2x)^2}$

take the square root of both sides: $0.33 = \frac{x}{0.33 - 2x}$
$0.11 - 0.66x = x$, then $1.66x = 0.11$ and $x = 0.066$
Therefore $[ICl] = 0.20 \, M$, $[I_2] = 0.066 \, M$ and $[Cl_2] = 0.066 \, M$.

24. $SO_2(g) + NO_2(g) \rightleftharpoons NO(g) + SO_3(g)$ $K_{eq} = 85.0$

**Table 19.3:**

<table>
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<th>0.100</th>
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<th>0</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>$-x$</td>
<td>$-x$</td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
<td>E</td>
<td>$0.100 - x$</td>
<td>$0.100 - x$</td>
<td>$x$</td>
<td>$x$</td>
</tr>
</tbody>
</table>

$K_{eq} = \frac{[NO][SO_3]}{[SO_2][NO_2]}$ or $85 = \frac{(x)(x)}{(0.100-x)(0.100-x)}$

take the square root of both sides: $9.22 = \frac{x}{0.100-x}$
then $0.922 = 10.22x$ and $x = 0.0902$

$[NO_2] = 9.8 \times 10^{-3} \, M$

$NO$

$= 0.0902 \, M$

Therefore the concentration of nitrogen dioxide is $0.8 \times 10^{-3} \, M$ and nitrogen monoxide is $0.0902 \, M$. 
19.2 Le Châtelier’s Principle

Lesson Quiz

Name___________________ Class______________ Date________

Multiple Choice:

1. Disruption in a chemical system in equilibrium can be caused by changes in
   1. catalyst
   2. time
   3. temperature
   4. reaction container

2. In the gas-phase reaction \(2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)\), an increase in pressure will
   1. favor the forward reaction
   2. favor the reverse reaction
   3. have no effect on the equilibrium
   4. decrease the rates of both reactions

3. In the reaction \(2NO + O_2 \rightleftharpoons 2NO_2\), an increase in \([O_2]\) will
   1. cause an increase in the forward reaction
   2. cause an increase in the reverse reaction
   3. cause a decrease in the forward reaction
   4. have no effect on concentrations

4. In the equilibrium reaction \(CaCO_3(s) \rightleftharpoons CaO(s) + O_2(g)\), an increase in pressure produces all of the following except
   1. an increase in CaCO\(_3\) formation
   2. an increase in the rate of the forward reaction
   3. a decrease in CaO formation
   4. an increase in the rate of the reverse reaction

5. In the reaction \(2SO_3 + \text{heat} \rightleftharpoons 2SO_2 + O_2\), an increase in temperature will produce
   1. an increase in the forward reaction
   2. an increase in the reverse reaction
   3. an increase in the formation of SO\(_3\)
   4. no effect on \(K_{eq}\)

6. In the gas phase reaction \(H_2 + Cl_2 \rightleftharpoons 2HCl\), an increase in volume will
   1. favor the forward reaction
   2. favor the reverse reaction
   3. have no effect on the equilibrium
   4. decrease the rates of both reactions

7. In the reaction \(HCl(aq) + CaCO_3(s) \rightleftharpoons CaCl_2(aq) + CO_2(g)\), addition of a catalyst will
   1. increase the \(K_{eq}\)
2. decrease the $K_{eq}$
3. have no effect on the $K_{eq}$
4. have variable effects on the $K_{eq}$

8. The reaction in problem seven will go to completion when
   1. $[\text{CaCO}_3]$ decreases
   2. $\text{CO}_2$ is removed from the reaction
   3. $[\text{CaCl}_2]$ increases
   4. $\text{CO}_2$ pressure increases

9. In the reaction $\text{CaCO}_3(s) \rightleftharpoons \text{CaO}(s) + \text{CO}_2(g)$, an increase in the amount of CaO will
   1. increase the $K_{eq}$
   2. decrease the $K_{eq}$
   3. have no effect on the $K_{eq}$
   4. have variable effects on the $K_{eq}$

10. In the reaction $\text{Ag}^+(aq) + \text{Cl}^-(aq) \rightleftharpoons \text{AgCl}(s)$, the reaction goes strongly in the forward direction because of
    1. formation of a precipitate
    2. formation of an ionized compound
    3. formation of a gas
    4. formation of water

**True/False:**

11. _____ If more $\text{H}_2$ is added to the Haber-Bosch process, the ammonia concentration will decrease.
12. _____ An increase in temperature in the Haber-Bosch process produces a decrease in ammonia formation.
13. _____ An increase in pressure on a solid stresses the system.
14. _____ In the Haber-Bosch process, an increase in volume increases the breakdown of ammonia.
15. _____ A change in pressure changes the $K_{eq}$ for a reversible reaction.
16. _____ The presence of a catalyst allows the equilibrium of a system to be reached more quickly.
17. _____ The formation of vapor-phase water causes a reaction to go to completion.
18. _____ A decrease in pressure causes gas molecules to collide more frequently
19. _____ A change in concentration of a reactant does not affect the value of $K_{eq}$.
20. _____ The $K_{eq}$ for an equilibrium process is the same at all temperatures.

**Fill in the blanks:**

21. When a chemical system that is at ________ is disturbed by a ______, the system will respond by attempting to ________ that stress until a new equilibrium is established.
22. If the ________ of one substance in a system is ________, the system will respond by favoring the reaction that ________ that substance.
23. An increase in the ________ of a system favors the direction of the reaction that absorbs heat, the ________ direction.
24. When the pressure is ________ by decreasing the available ________, the reaction that produces ________ total moles of gas becomes favored.
25. When one of the products of a reaction is ________ from the chemical equilibrium system as soon as it is ________, the ________ reaction cannot establish itself, and ________ is never reached
26. List the situations in which an equilibrium reaction can go to completion. Give an example for each situation.

27. Blood levels of hydrogen ions are regulated by the equilibrium process:
28. $\text{H}_2\text{O}(l) + \text{CO}_2(g) \rightleftharpoons \text{H}^+(aq) + \text{HCO}_3^-(aq)$
a. In a panic attack, the individual has short, shallow respirations leading to a decrease in blood levels of CO2. Which way does the equilibrium shift? What is the effect on [H+]?

b. Some clinical conditions produce decreased respiration and elevated blood CO2 concentrations. Which way does the equilibrium shift? What is the effect on [H+]?

29. Consider the following reaction:

\[ \text{CO}_2(g) + \text{H}_2(g) \rightleftharpoons \text{CO}(g) + \text{H}_2\text{O}(g) \]

30. Calculate the value of the equilibrium constant, \( K_{eq} \), for the above system, if 0.1908 moles of CO, 0.0908 moles of H, 0.0092 moles of CO, and 0.0092 moles of H2O vapor were present in a 2.00 L reaction vessel at equilibrium.

40. Consider the following reaction:

\[ \text{C}_2\text{H}_4(g) + \text{H}_2(g) \rightleftharpoons \text{C}_2\text{H}_6(g) \]

42. With a \( K_{eq} = 0.99 \). What is the concentration for each substance at equilibrium if the initial concentration of ethene, C, H, and H2 are 0.335 M and that of hydrogen is 0.526 M?

---

**Answer Key**

1. c
2. b
3. a
4. b
5. a
6. c
7. c
8. b
9. c
10. a
11. false
12. true
13. false
14. true
15. false
16. true
17. false
18. false
19. true
20. false
21. equilibrium, stress, counteract
22. concentration, increased, removes
23. temperature, endothermic
24. increased, volume, fewer
25. removed, produced, reverse, equilibrium

1. Formation of a precipitate: \(AgNO_3(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + AgCl(s)\)
2. Formation of a gas: \(Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)\)
3. Formation of water: \(HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)\)

1. Equilibrium shifts to the left and \([H^+]\) decreases.
2. Equilibrium shifts to the right and \([H^+]\) increases.

26. \([CO_2] = 0.1908 \text{ mol} \quad \frac{CO_2}{L} = 0.0954 \text{ M}

27. \([H_2] = 0.0454 \text{ M}\)

CO

\[= 0.0046 \text{ M}\]

\(H_2O\)

\[= 0.0046 \text{ M}\]

28. \(K_{eq} = \frac{(0.0046)(0.0046)}{(0.0954)(0.0954)} = 4.9 \times 10^{-3}\)

**Table 19.4:**

<table>
<thead>
<tr>
<th>Material</th>
<th>(C_2H_4)</th>
<th>(H_2)</th>
<th>(C_2H_6)</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>0.335</td>
<td>0.526</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>(−x)</td>
<td>(−x)</td>
<td>(+x)</td>
</tr>
<tr>
<td>E</td>
<td>0.335 − (x)</td>
<td>0.526 − (x)</td>
<td>(+x)</td>
</tr>
</tbody>
</table>

\(K_{eq} = \frac{x}{(0.335−x)(0.526−x)} = 0.0995\) use that pesky quadratic equation again

\([C_2H_4] = 0.236 \text{ M} \quad [H_2] = 0.427 \text{ M} [C_2H_6] = 0.0995 \text{ M}\)
Lesson Quiz

Name___________________ Class______________ Date________

Multiple Choice:

1. All of the following are true of saturated solutions except
   1. a saturated solution has the maximum amount of solute dissolved
   2. has undissolved solute present
   3. can dissolve more solute
   4. has an equilibrium between dissolved and undissolved solute

2. The $K_{sp}$ for $\text{PbI}_2$ is written
   1. $[\text{Pb}]^2[I]$
   2. $\frac{[\text{Pb}]}{[I]^2}$
   3. $[\text{Pb}][I]^2$
   4. $\frac{[\text{Pb}]}{[I]^2}$

3. For the reaction $\text{AgBr}(s) \rightleftharpoons \text{Ag}^+(aq) + \text{Br}^-(aq)$, the $K_{sp}$ is
   1. $[\text{Ag}][\text{Br}]$
   2. $\frac{[\text{Ag}][\text{Br}]}{[\text{AgBr}]}$
   3. $\frac{[\text{Ag}][\text{Br}]}{[\text{Ag}][\text{Br}]}$
   4. $\frac{[\text{Ag}]}{[\text{Br}]}$

4. The units for $K_{sp}$ are
   1. moles/L
   2. grams/L
   3. unitless
   4. depends on the expression

5. When two solutions are mixed that could form a precipitate, the precipitate will occur when
   1. the ion product is greater than the $K_{sp}$.
   2. the ion product equals the $K_{sp}$.
   3. the ion product is less than the $K_{sp}$.
   4. the $K_{sp}$ is greater than the ion product.

6. If sodium bicarbonate is added to a solution of calcium carbonate in equilibrium with its ions, $\text{CaCO}_3(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq)$ the added material will
   1. increase the solubility of calcium carbonate
   2. decrease the solubility of calcium carbonate
   3. have no effect on the solubility of calcium carbonate
   4. change the $K_{sp}$ for the process

True/False:
7. _____ KBr is highly soluble in water.
8. _____ CaSO₄ is less soluble that Ca₃(PO₄)₂.
9. _____ Molarity is used to convert between solubility and molar solubility.
10. _____ Calculating the ion product for a material allows prediction of precipitate formation.
11. _____ The K_sp for a compound can be calculated if the molar solubility is known.
12. _____ Most insoluble ionic compounds will still dissolve to a small extent.

Fill in the blanks:

13. The solubility product constant, (K_sp) is equal to the mathematical _______ of the ions, each raised to the power of the _______ of the ion in the _______ equation.
14. Molar _______ is the number of moles of _______ in one liter of a _______ solution.
15. If the value of the ion _______ is greater than the value of K_sp, then a _______ will form.
16. A _______ ion is an ion that is common to more than one _______ in a solution.
17. The common ion effect is a _______ in the _______ of an _______ compound as a result of the addition of a common ion.
18. The solubility of barium sulfate at 298 K is 1.05 × 10⁻⁵ M. Calculate the K_sp.
19. The solubility of Mg(OH)₂ is 1.71 × 10⁻⁴ M. Calculate the K_sp.
20. The K_sp for AgCl is 1.8 × 10⁻¹⁰. What is the molar solubility of AgCl in pure water?
21. How much SrF₂(K_sp = 2.5 × 10⁻⁹) will dissolve in one liter of water?
22. Determine if a precipitate of Ca(OH)₂ will form if 20 mL of 0.30 M CaCl₂ solution is mixed with 10 mL of 0.90 M NaOH. The K_sp for Ca(OH)₂ is 5.1 × 10⁻₆.
23. AgCl will be dissolved into a solution with is already 0.0100 M in chloride ion. What is the solubility of AgCl?

Answer Key

1. c
2. c
3. a
4. d
5. a
6. b
7. true
8. false
9. false
10. true
11. true
12. true
13. product, coefficient, dissociation
14. solubility, solute, saturated
15. product, precipitate
16. common, salt
17. decrease, solubility, ionic
18. \( BaSO₄ \rightleftharpoons Ba^{2+} + SO_{4}^{2-} \)
19. \( K_{sp} = [Ba^{2+}] [SO_{4}^{2-}] \)
20. so
21. \([Ba^{2+}] = [SO_{4}^{2-}] \)
22. then
23. \( K_{sp} = (1.05 \times 10^{-5} M)^2 = 1.10 \times 10^{-10} \)
24. \( Mg(OH)_2 \rightleftharpoons Mg^{2+} + 2OH^- \)
25. \( K_{sp} = [Mg^{2+}][OH^-]^2 \quad [Mg^{2+}] = 1.71 \times 10^{-4} M \)
26. and
27. \( [OH^-] = 2(1.71 \times 10^{-4} M) \)
28. \( K_{sp} = (1.71 \times 10^{-4})(3.42 \times 10^{-4})^2 = 2.00 \times 10^{-11} \)
29. Let \( x \) be the molar solubility, then
30. 

\[
AgCl = Ag^+ + Cl^- \\
K_{sp} = [Ag][Cl] = x^2 \\
x^2 = 1.8 \times 10^{-10} \\
x = 1.34 \times 10^{-5} M
\]
31. \( SrF_2 \rightleftharpoons Sr^{2+} + 2F^- \)
32. at equilibrium,
33. \([Sr^{2+}] = x\)
34. and
35. \([F^-] = 2x\)
36. \(K_{sp} = [Sr^{2+}][F^-]^2\)
37. or
38. \(2.5 \times 10^{-9} = (x)(2x)^2\)
39. \(4x^3 = 2.5 \times 10^{-9}\)
40. or
41. \(x^3 = 6.25 \times 10^{-10}\)
42. and
43. \(x = 8.5 \times 10^{-4} M\)
44. \(8.5 \times 10^{-4}\)
45. moles of SrF
46. 2
47. will dissolve in one liter of water.
48. \(Ca(OH)_2 \rightleftharpoons Ca^{2+} + 2OH^-\) and \(K_{sp} = [Ca^{2+}][OH^-]^2\)
49. need \([Ca]^{2+}\)
50. 2
51. ] and \([OH^-]\)
52. 
53. ] total volume of mixture = 30 mL for \([Ca]^{2+}\)
54. 2
55. ]
56. \(V_1C_1 = V_2C_2\)
57. or
58. \(C_2 = \frac{V_1C_1}{V_2}\)
59. and
60. \(C_2 = (0.02 \, L)(\frac{0.3 \, moles/L}{0.03 \, L}) = 0.2 \, M\)
61. for \([OH^-]\)
62. 
63. ]
64. \(V_1C_1 = V_2C_2\)
65. or
66. \(C_2 = \frac{V_1C_1}{V_2}\)
67. and
68. \( C_2 = \frac{(0.01 \text{ L})(0.9 \text{ moles/L})}{0.03 \text{ L}} = 0.3 \text{ M} \)

69. ion product for

70. \( Ca(OH)_2 = (0.2)(0.3)^2 = 0.018 \)

71. value is larger than \( K_{sp} \)

72. \( s_p \)

73. , so a precipitate will form.

74. \( K_{sp} = [Ag^+][Cl^-] \) so \( 1.77 \times 10^{-10} = [Ag^+][Cl^-] \)

75. let

76. \( x = [Ag^+] \)

77. and

78. \( 0.0100 + x = [Cl^-] \)

79. and assume

80. \( x \)

81. is small enough that [Cl ] is essentially 0.0100 M then

82. 

83. \( 1.77 \times 10^{-10} = (x)(0.0100) \)

84. and

85. \( x = 1.77 \times 10^{-8} \text{ M} \)

86. concentration of Ag

87. +
19.4 Equilibrium

Chapter Test

Name___________________ Class______________ Date________

Multiple Choice:

1. A reversible reaction is indicated by
   1. $\rightarrow$
   2. $\leftrightarrow$
   3. $\leftrightarrow$
   4. $\leftrightarrow$

2. In the reversible reaction between CaCO$_3$(s) and its products CaO(s) and O$_2$(g), the equilibrium expression is written
   1. $[CaO][O_2]$
   2. $[O_2]$
   3. $\frac{[CaO][O_2]}{[CaCO_3]}$
   4. $\frac{[CaO]}{[CaCO_3]}$

3. In the equilibrium gas-phase reaction 2NOCl(g) produces 2NO(g) + Cl$_2$(g), a decrease in pressure will
   1. favor the forward reaction
   2. favor the reverse reaction
   3. have no effect on the equilibrium
   4. decrease the rates of both reactions

4. In the equilibrium gas phase reaction $H_2 + Cl_2$ forms 2 HCl, a decrease in volume will
   1. favor the forward reaction
   2. favor the reverse reaction
   3. have no effect on the equilibrium
   4. decrease the rates of both reactions

5. If lead nitrate is added to a solution of calcium carbonate in equilibrium with its ions Ca$^{2+}$(aq) + CO$_3^{2-}$(aq), the added lead will
   1. increase the formation of calcium carbonate
   2. decrease the formation of calcium carbonate
   3. have no effect on the formation of calcium carbonate
   4. change the K$_{sp}$ for the process

6. When two solutions are mixed that could form a precipitate, no precipitate will form unless
   1. the ion product is greater than the K$_{sp}$.
   2. the ion product equals the K$_{sp}$
   3. the ion product is less than the K$_{sp}$
   4. the K$_{sp}$ is greater than the ion product

7. A value of K$_{eq}$ less than 1 means
19.4. Equilibrium

1. products favored over reactants
2. reactants favored over products
3. an equilibrium situation
4. reaction has not yet come to completion

8. In the reaction $2SO_3 + \text{heat} \rightleftharpoons 2SO_2 + O_2$, a decrease in temperature will produce
   1. an increase in the forward reaction
   2. an increase in the reverse reaction
   3. an increase in the formation of SO$_3$
   4. no effect on $K_{eq}$

9. In the Haber-Bosch reaction, if additional N$_2$ is added to the system, at equilibrium all of the following will be seen except
   1. the final concentration of N$_2$ will be higher
   2. the final concentration of H$_2$ will be higher
   3. the final concentration of NH$_3$ will be higher
   4. the final concentration of H$_2$ will be lower

10. A decrease in temperature in the Haber-Bosch process results in all of the following except
    1. an increase in [NH$_3$]
    2. a decrease in $K_{eq}$
    3. a decrease in [N$_2$]
    4. an increase in $K_{eq}$

True/False:

11. _____ All equilibrium reactions are reversible.
12. _____ The $K_{eq}$ does not depend on the direction of the equilibrium reaction.
13. _____ The formation of liquid water causes a reaction to go to completion.
14. _____ A change in concentration of a reactant changes the value of $K_{eq}$
15. _____ A change in pressure on a liquid has no effect on $K_{eq}$.
16. _____ Pb(NO$_3$)$_2$ is very soluble in water.
17. _____ The presence of a catalyst has no effect on how fast the equilibrium of a system is reached.
18. _____ A decrease in volume causes gas molecules to collide more frequently.
19. _____ When the concentrations of reactants and products have become equal, equilibrium has been achieved.
20. _____ The equilibrium position is the property of the particular reversible reaction.

Fill in the blanks:

21. Define the following terms:
   1. reversible reaction
   2. chemical equilibrium
   3. solubility product constant
   4. Le Châtelier’s principle

22. List the conditions and properties of a system at equilibrium.
23. The equilibrium position for a given reaction does not depend on the _______concentrations, so the equilibrium constant has the same _______ regardless of the _______ amounts of each reaction component.
24. Rank the following compounds in terms of solubility, beginning with the most soluble and ending with the least soluble:
## Table 19.5:

<table>
<thead>
<tr>
<th>Compound</th>
<th>$K_{eq}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>CaCO$_3$</td>
<td>$4.5 \times 10^{-9}$</td>
</tr>
<tr>
<td>CuS</td>
<td>$8.0 \times 10^{-3}$</td>
</tr>
<tr>
<td>PbSO$_4$</td>
<td>$6.3 \times 10^{-7}$</td>
</tr>
<tr>
<td>CaSO$_4$</td>
<td>$2.4 \times 10^{-3}$</td>
</tr>
</tbody>
</table>

25. Write the $K_{eq}$ expressions for the following compounds

1. Al(OH)$_3$
2. AgCl
3. Ca$_3$(PO$_4$)$_2$
4. BaCO$_3$

26. Calculate the $K_{eq}$ for the following gas-phase reaction:

27. $CO_2 + H_2 \rightleftharpoons CO + H_2O$

28. $[CO_2] = 0.095 \ M \ [H_2] = 0.045 \ M \ [CO] = 0.0046 \ M \ [H_2O] = 0.0046 \ M$

29. If one requirement for an equilibrium system is that it is closed, what would be the eventual fate of CaCO$_3$ if left exposed to the outside air?

30. At 448°C, the $K_{eq}$ for the reaction $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ is 50.0. In which direction is the reaction most likely to go? How many moles of HI will be present at equilibrium when 1.0 mole $H_2$ and 1.0 mole $I_2$ react in a 1.0 L container?

31. What is the concentration of the silver ion in 1.0 L of a saturated solution of AgCl if 0.05 moles of NaCl are added to the solution?

### Answer Key

1. c
2. b
3. a
4. c
5. b
6. a
7. b
8. b
9. d
10. b
11. true
12. false
13. true
14. false
15. true
16. true
17. false
18. true
19. false
20. true

1. A reversible reaction is a reaction in which the conversion of reactants to products and the conversion of products to reactants occur simultaneously.
2. Chemical equilibrium is the state of a system in which the rate of the forward reaction is equal to the rate of the reverse reaction.
3. The solubility product constant, \( K_{sp} \), is equal to the mathematical product of the ions, each raised to the power of the coefficient of the ion in the dissociation equation.
4. When a chemical system that is at equilibrium is disturbed by a stress, the system will respond by attempting to counteract that stress until a new equilibrium is established.

1. The system must be closed, meaning no substances can enter or leave the system.
2. Equilibrium is a dynamic process. Even though we don’t observe any changes, both the forward and reverse reactions are still taking place.
3. The rates of the forward and reverse reactions must be equal.
4. The amounts of reactants and products do not have to be equal. However, after equilibrium is attained, the amounts of reactants and products will remain constant.

21. starting, value, initial
22. most soluble CaSO\(_4\)
23. PbSO\(_4\)
24. CaCO\(_3\)
25. least soluble CuS

27. least soluble CuS

1. \( K_{eq} = [Al][OH]^3 \)
2. \( K_{eq} = [Ag][Cl] \)
3. \( K_{eq} = [Ca]^3[PO_4]^4 \)
4. \( K_{eq} = [Ba][CO_3] \)

28. \( K_{eq} = \frac{[CO][H_2]O}{[CO_2][H_2]} = \frac{(0.0046)(0.0046)}{(0.095)(0.045)} = 4.9 \times 10^{-3} \)

29. The equilibrium reaction involves the reversible formation of CaO and CO\(_2\). If exposed, the CO\(_2\) will leave the site and the calcium carbonate will eventually disappear.

30. With a large and positive \( K_{eq} \), the reaction will go strongly toward the formation of HI.

**Table 19.6:**

<table>
<thead>
<tr>
<th></th>
<th>([H_2])</th>
<th>([I_2])</th>
<th>([HI])</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>1.0</td>
<td>1.0</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>(-x)</td>
<td>(-x)</td>
<td>(+2x)</td>
</tr>
<tr>
<td>E</td>
<td>(1.0 - x)</td>
<td>(1.0 - x)</td>
<td>(2x)</td>
</tr>
</tbody>
</table>

\[ K_{eq} = \frac{[HI]^2}{[H_2][I_2]} \] then 50 = \( \frac{(2x)^2}{(1.0-x)^2} \) and \( \sqrt{50} = \frac{2x}{1.0-x} \)

7.07 - 7.07x = 2x, then 7.07 = 9.07x and x = 0.78 at equilibrium, there are 2 x 0.78 = 1.6 *moles* HI present

29. \( K_{sp} \) for AgCl = \( 1.8 \times 10^{-10} \)
30. let
31. \( x = \) concentration of \( Ag^+ \)
32. , then
33. \( [Cl^-] = 0.05 + x \)
34. then assuming
35. \( x \)
36. is very small,
37. \( [Cl^-] = 0.05 \)
38. we can write
39. \( K_{sp} = [x][0.05] \)
40. and
41. \( x = 1.8 \times 10^{-10} \div 0.05 = 3.6 \times 10^{-9} \, M[Ag^+] \)
Chapter 20

Entropy and Free Energy Assessments

Chapter Outline

20.1 Entropy
20.2 Spontaneous Reactions and Free Energy
20.3 Free Energy and Equilibrium
20.4 Entropy and Free Energy
20.1 Entropy

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. One of the following is a characteristic of an exothermic reaction
   1. products have high energy as compared to reactants
   2. heat is absorbed during the reaction
   3. products are more stable than reactants
   4. energy is needed to drive the reaction

2. One of the following is a characteristic of an endothermic reaction
   1. energy is absorbed during the reaction
   2. reaction is energetically favorable
   3. products have lower quantity of energy than reactants
   4. products are more stable than reactants

3. An increase in entropy is seen in one of the following situations
   1. metallic Na melts
   2. CO₂ gas condenses
   3. formation of AgCl precipitate
   4. cleaning up your room

4. A decrease in entropy is seen in one of the following situations
   1. \( CaCO₃ \rightleftharpoons CaO + CO₂ \)
   2. \( H₂O(l) \rightarrow H₂O(g) \)
   3. crystallization of sucrose
   4. \( 2HgO \rightarrow 2Hg + O₂ \)

5. One of the following illustrates a decrease in entropy
   1. increase in temperature
   2. decrease in temperature
   3. increase in kinetic energy
   4. increase in motion of particles

6. \( S \) is the symbol for
   1. enthalpy
   2. energy
   3. entropy
   4. epitome

7. The \( S° \) for ice would have an entropy value
   1. higher than that for liquid water
   2. lower than that for liquid water
3. unable to predict
4. higher than that for steam

8. The standard entropy for diamond is less than that for graphite because
   1. graphite is more highly organized than diamond
   2. the two materials are composed of different elements
   3. diamond is more highly organized than graphite
   4. two different chemical reactions are involved

9. All molecular motion ceases at
   1. 100 K
   2. -250 K
   3. 272 K
   4. -272 K

10. The melting of ice involves all of the following except
    1. decrease in temperature
    2. change of state
    3. increase in entropy
    4. breaking of hydrogen bonds

True/False:

11. _____ The standard entropy of a substance is given by ΔS.
12. _____ The entropy change for the sublimation of CO₂ is positive.
13. _____ Most naturally occurring reactions are endothermic.
14. _____ The energy of the system decreases during an exothermic reaction.
15. _____ Work must be done to increase entropy.
16. _____ The entropy of a gas is less than the entropy of the corresponding liquid.
17. _____ The solid state of MgSO₄ is highly ordered.
18. _____ Endothermic reactions can occur spontaneously.
19. _____ The process of dissolving decreases entropy.
20. _____ The entropy of a material at absolute zero is negative.

Fill in the Blank:

21. In an exothermic reaction, the _______ have a relatively high quantity of ______ compared to the ______-_____.

22. _______ is a measure of the degree of randomness or ______ of a system.

23. An increase in ______ means that the ______ of the substance have greater ______ energy

24. In each of the following cases, indicate whether the process results in an increase or decrease in the entropy of the system and explain with a brief comment.
   1. baking a cake
   2. setting off fireworks
   3. photosynthesis: 6CO₂(g) + 6H₂O(l) + light → C₆H₁₂O₆(aq) + 6O₂(g)
   4. burning of propane: C₃H₈(g) + 5O₂(g) → 4H₂O(g) + 3CO₂(g) + energy
   5. rusting of iron: Fe(s) + O₂(g) + H₂O(l) → Fe₂O₃·nH₂O(s)
   6. vinegar/baking soda: HC₂H₃O₂(aq) + NaHCO₃(aq) → NaC₂H₃O₂(aq) + H₂O(l) + CO₂(g)

25. Calculate the standard entropy change for the following reaction:
   26. Cu(s) + ½O₂(g) → CuO(s)
   27. given that
   28. S°[Cu(s)] = 33.15 J/K⋅mol
29. 
30. \( S^\circ [O_2(g)] = 205.14 \, J/K \cdot mol \)
31. 
32. \( S^\circ [CuO(s)] = 42.63 \, J/K \cdot mol \)
33. 
34. Calculate \( \Delta S^\circ \) for the following reaction:
35. \( H_2(g) + Cl_2(g) \rightarrow 2HCl(g) \)
36. \( \Delta S^\circ [H_2] = 131.0 \, J/K \cdot mol \)
37. 
38. \( \Delta S^\circ [Cl_2] = 223.0 \, J/K \cdot mol \)
39. 
40. \( \Delta S^\circ [HCl] = 187.0 \, J/K \cdot mol \)
41. 
42. Calculate \( \Delta S^\circ \) for the reaction: \( N_2O(g) \rightarrow N_2(g) + \frac{1}{2}O_2(g) \)
43. \( \Delta S^\circ \)
44. for
45. \( N_2O(g) = 220.0 \, J/K \cdot mol \)
46. 
47. \( \Delta S^\circ \)
48. for
49. \( N_2(g) = 192.0 \, J/K \cdot mol \)
50. 
51. \( \Delta S^\circ \)
52. for
53. \( O_2(g) = 205.14 \, J/K \cdot mol \)
54. 

Answer Key

1. c
2. a
3. a
4. c
5. b
6. c
7. b
8. c
9. d
10. a
11. false
12. true
13. false
14. false
15. false
16. false
17. true
18. true
19. false
20. false
21. reactants, energy, products
22. entropy, disorder
23. temperature, particles, kinetic
   1. decrease, formation of solid from liquids
   2. increase, combustion and formation of gases from solids
   3. decrease, fewer particles in products than in reactants
   4. increase, release of energy plus formation of gas and increase in number of particles
   5. decrease, formation of solid, conversion of gas and liquid reactants to solid product
   6. increase, more product particles than reactant particles and formation of gas

24. $\Delta S^\circ = 42.63 - [33.15 + \frac{1}{2}(205.14)] = -93.09 \text{ J/K} \cdot \text{mol}$
25. $\Delta S^\circ = (2)(187.0 \text{ J/K} \cdot \text{mol}) - [(131.0 \text{ J/K} \cdot \text{mol}) + 223.0 \text{ J/K} \cdot \text{mol}] = 20.0 \text{ J/K} \cdot \text{mol}$
26. $\Delta S^\circ = [192.0 + \frac{1}{2}(205.14)] - 220.0 = 74.6 \text{ J/K} \cdot \text{mol}$
20.2 Spontaneous Reactions and Free Energy

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. A spontaneous reaction has
   1. an increase in enthalpy
   2. an increase in entropy
   3. absorption of energy
   4. decrease in randomness

2. In a nonspontaneous reaction we observe
   1. an increase in entropy
   2. randomness increases
   3. endothermic reaction
   4. products are favored over reactants

3. The dissolving of AgNO₃ in water is a spontaneous reaction because all of the following except
   1. a precipitate is formed
   2. entropy is increased
   3. randomness is increased
   4. there is a change of state

4. If ΔH is positive and ΔS is positive, then ΔG is
   1. always negative
   2. never negative
   3. negative at higher temperatures
   4. negative at lower temperatures

5. If ΔH is negative and ΔS is negative, then ΔG is
   1. always negative
   2. never negative
   3. negative at higher temperatures
   4. negative at lower temperatures

6. When the TΔS term is numerically smaller than the ΔH term, the ΔG value will be
   1. positive if ΔS is negative
   2. negative if ΔS is positive
   3. negative if ΔS is negative
   4. variable value

7. To be used in Gibbs free energy calculations, ΔS values
   1. can be used without conversion
   2. must be converted to kJ/mol
3. must be converted to J/mol
4. must be divided by 100

8. Energy that is available to do work is called
   1. stored energy
   2. entropic energy
   3. free energy
   4. enthalpy

9. The reaction \( CO_2(g) + 2H_2O(g) \rightarrow \Delta CH_4(g) + 2O_2(g) \) is nonspontaneous because
   1. entropy is increased
   2. reaction is endothermic
   3. reaction is exothermic
   4. reaction involves change of state

10. The dissolving of KCl in water is spontaneous because of all the following except
    1. entropy is increased
    2. solid is converted to aqueous particles
    3. no energy needed for process to occur
    4. reactants favored over products

True/False:

11. _____ The combustion of wood is spontaneous at room temperature.
12. _____ Entropy increases during most combustion reactions.
13. _____ \( \Delta G \) is positive for a spontaneous reaction.
14. _____ An endothermic reaction has a positive \( \Delta H \).
15. _____ \( \Delta G \) is positive above 273 K for the freezing of water.
16. _____ Standard conditions for measuring entropy and enthalpy values are one atm pressure and 298 K temperature.
17. _____ The products of a combustion reaction are primarily carbon dioxide gas and water vapor.
18. _____ Lightning can cause the formation of NO.
19. _____ Conversion of liquid water to steam is spontaneous above 300 K.
20. _____ Release of \( CO_2 \) from an opened cold soft drink is a spontaneous process.

Fill in the Blank:

21. A spontaneous reaction is a reaction that favors the ________ of products at the conditions under which the reaction is occurring.
22. The_______ of a fire are composed mostly of gases such as carbon dioxide and ____________, so the ________ of the system increases during most combustion reactions.
23. A ________ reaction is a reaction that does not favor the formation of ________at the given set of conditions.
24. Write the mathematical equation for the calculation of Gibbs free energy.
25. When one driving force favors the _______, but the other does not, it is the ________ that determines the sign of \( \Delta G \).
26. Which of the following is spontaneous?
   1. \( \Delta H > 0, T(\Delta S) > 0 \)
   2. \( \Delta H < 0, T(\Delta S) > 0 \)
   3. \( \Delta H > 0, T(\Delta S) < 0 \)
   4. \( \Delta H < 0, T(\Delta S) < 0 \)

Short Answer:
27. For the reaction \( Cu_2S(s) + 2S(s) \rightarrow 2CuS(s) \)
28. \( \Delta H^\circ = -26.7 \text{ kJ/mol} \)
29. and
30. \( \Delta S^\circ = -19.7 \text{ J/K/mol} \)
31. under standard conditions Calculate
32. \( \Delta G^\circ \)
33. .
34. For a certain process at 300.0 K, \( \Delta G = -77.0 \text{ kJ/mol} \) and \( \Delta H = -56.9 \text{ kJ/mol} \). Find the entropy change for this process.
35. The hydrogenation of ethene gas under standard conditions \((T = 298.15 \text{ K})\) shows a decrease in disorder \( (\Delta S^\circ = -0.1207 \text{ kJ/(K\cdot mol)}) \) during an exothermic reaction \( (\Delta H^\circ = -136.9 \text{ kJ/mol}) \). Calculate \( \Delta G^\circ \).
36. The entropy of a system at 337.1 K increases by 221.7 \text{ J/K/mol}. The free energy value is found to be -717.5 \text{ kJ/mol}. Calculate the change in enthalpy of this system.

**Answer Key**

1. b
2. c
3. a
4. c
5. d
6. a
7. b
8. c
9. b
10. d
11. false
12. true
13. false
14. true
15. true
16. true
17. true
18. true
19. false
20. true
21. formation
22. products, water vapor, entropy
23. nonspontaneous, products
24. \( \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ \)
25. reaction, temperature
26. b
27. \( \Delta S = -0.0197 \text{ kJ/K\cdot mol} \)
28. \( \Delta G^\circ = -26.7 \text{ kJ/mol} - (298)(-0.0197 \text{ kJ/K\cdot mol}) = -26.7 + 5.9 = 20.8 \text{ kJ/mol} \)
29. \( \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ \), then \(-T\Delta S^\circ = \Delta G^\circ - \Delta H^\circ \) or \( T\Delta S^\circ = \Delta H^\circ - \Delta G^\circ \)
30. \( \Delta S = \frac{\Delta H - \Delta G}{T} = \frac{(-56.9 \text{ kJ/mol}) - (-77.0 \text{ kJ/mol})}{398 \text{ K}} = 0.067 \text{ kJ/K\cdot mol} = 67 \text{ J/K\cdot mol} \)
31. \( \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = -136.9 \text{ kJ/mol} - (-0.1207 \text{ kJ/K\cdot mol})(298.15 \text{ K}) = 100.9 \text{ kJ/mol} \)
32. \( \Delta S^\circ = 221.7 \text{ J/K\cdot mol} = 0.2217 \text{ kJ/K\cdot mol} \)
33. \( \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ \) so \( \Delta H^\circ = \Delta G^\circ + T\Delta S^\circ = -717.5 \text{ kJ/mol} + (337.1 \text{ K})(0.2217 \text{ kJ/K\cdot mol}) = -642.9 \text{ kJ/mol} \)
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. The quicklime manufacturing process has all the following characteristics except
   1. reaction favors reactants at lower temperatures
   2. reaction becomes spontaneous above 835°C
   3. CO₂ is trapped in reaction vessel
   4. as CO₂ pressure approaches one atm, products are favored over reactants

2. In the quicklime manufacturing process, the amount of products is detected by
   1. measuring disappearance of CaCO₃
   2. measuring CO₂ pressure
   3. measuring CaO formation
   4. measuring T

3. When ∆G for a reversible reaction = zero,
   1. the reaction becomes endothermic
   2. the reaction becomes exothermic
   3. ∆S = zero
   4. ∆H = zero

4. At the transition from liquid water to steam
   1. ∆G for the process is positive
   2. ∆G for the process is negative
   3. ∆H for the process is zero
   4. ∆S for the process is positive

5. The value for R is
   1. 8.314 J/K·mol
   2. 8.314 kJ/K·mol
   3. 8.314 J/mol
   4. 8.314 kJ/mol

6. When Kₑq is less than one, all of the following are true except
   1. ∆G is positive
   2. ln(Kₑq) is negative
   3. ln(Kₑq) is positive
   4. reactants are favored

7. As the solid state changes into the liquid state
   1. ∆S becomes positive
   2. ∆S becomes negative
3. \( \Delta S = 0 \)
4. \( \Delta H_{fus} \) changes

8. When calculating \( K_{eq} \), units for \( \Delta G \) need to be
   1. \( \text{kJ/mol} \)
   2. \( \text{kJ/K} \cdot \text{mol} \)
   3. \( \text{J/K} \cdot \text{mol} \)
   4. \( \text{J/mol} \)

True/False:

9. _____ When a system is at equilibrium, neither the forward or reverse reactions are spontaneous.
10. _____ At room temperature, the quicklime reaction has a large positive \( \Delta G \).
11. _____ The quicklime process produces measurable amounts of \( \text{CO}_2 \) at room temperature.
12. _____ \( \Delta S_{fus} \) represents the entropy change during the melting process.
13. _____ When \( K_{eq} = 1 \), \( \ln(K_{eq}) = 1 \)
14. _____ We can calculate \( K_{eq} \) if we know \( \Delta G \).
15. _____ A large positive free energy change produces a very large \( K_{eq} \).
16. _____ \( K_{sp} \) can be used to calculate \( \Delta G \).

Fill in the Blank:

17. Carbon tetrachloride has a melting point of 250 K and the \( \Delta H_{fus} = 2.67 \text{kJ/mol} \). Calculate the \( \Delta S_{fus} \).
18. Nonane (C\(_9\)H\(_{20}\)) boils at 491 K and the \( \Delta H_{vap} = 40.5 \text{kJ/mol} \). Calculate the \( \Delta S_{vap} \).
19. For the following reaction at 327 C.
   \[
   N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)
   \]
   20. given
   21. \( \Delta H = -91.8 \text{kJ/mol} \)
   22. and
   23. \( \Delta S = -197.3 \text{J/K} \cdot \text{mol} \)
   24. Calculate \( \Delta G \)
   25. Calculate the temperature where the system will be at equilibrium

26. For the reaction at 400 K:
27. \( \text{CoCl}_2(g) \rightarrow CO(g) + Cl_2(g) \)
28. given
29. \( \Delta H = 110.5 \text{kJ/mol} \)
30. and
31. \( \Delta S = 136.8 \text{J/K} \cdot \text{mol} \)
   26. Calculate \( \Delta G \)
   27. Calculate the temperature at which the reaction is at equilibrium.

32. For the reaction \( \text{BaSO}_4(s) \rightleftharpoons \text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \), the \( K_{sp} = 1.1 \times 10^{-10} \) at 298 K.
33. Calculate the
34. \( \Delta G \)
35. for this process.
36. For the reaction \( \text{CO}_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g) \) at 298 K, \( \Delta G = 28.9 \text{kJ/mol} \)
37. Calculate the
38. \( K_{eq} \)
39. for the reaction.
Answer Key

1. c
2. b
3. b
4. d
5. a
6. b
7. a
8. d
9. true
10. true
11. false
12. true
13. false
14. true
15. false
16. true
17. \(0 = \Delta H_{\text{fus}} - T \Delta S_{\text{fus}}\) so \(\Delta S_{\text{fus}} = \frac{2.67 \text{ kJ/mol}}{250 \text{ K}} = \frac{2670 \text{ J/mol}}{250 \text{ K}} = 10.7 \text{ J/K} \cdot \text{mol}\)
18. \(\Delta S_{\text{vap}} = \frac{\Delta H_{\text{vap}}}{T} = \frac{40.5 \text{ kJ/mol}}{491 \text{ K}} = \frac{40500 \text{ J/mol}}{491 \text{ K}} = 82.5 \text{ J/K} \cdot \text{mol}\)
19. \(\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ\) so \(\Delta G = -91.8 \text{ kJ/mol} - (600 \text{ K})(-0.197 \text{ kJ/K} \cdot \text{mol}) = -91.8 \text{ kJ/mol} + 118.2 \text{ kJ/mol}\) so \(\Delta G = +26.4 \text{ kJ/mol}\)
   1. at equilibrium \(0 = \Delta H^\circ - T \Delta S^\circ\) then \(T = \frac{\Delta H}{\Delta S} = \frac{-91.8 \text{ kJ/mol}}{-0.197 \text{ kJ/K} \cdot \text{mol}} = 466 \text{ K}\)
20. \(\Delta S = 136.8 \text{ J/K} \cdot \text{mol} = 0.1368 \text{ kJ/K} \cdot \text{mol}\)
   1. \(\Delta G = 110.5 \text{ kJ/mol} - (400 \text{ K})(0.1368 \text{ kJ/K} \cdot \text{mol}) = 110.5 \text{ kJ/mol} - 54.5 \text{ kJ/mol} = 56.0 \text{ kJ/mol}\)
   2. \(\Delta T = \frac{110.5 \text{ kJ/mol}}{0.1368 \text{ kJ/K} \cdot \text{mol}} = 808 \text{ K}\)
21. \(\Delta S = -46.4 \text{ J/K} \cdot \text{mol} = -0.0464 \text{ kJ/K} \cdot \text{mol}\)
22. \(\Delta G = -34.90 \text{ kJ/mol} - (298 \text{ K})(-0.0464 \text{ kJ/K} \cdot \text{mol}) = -34.90 \text{ kJ/mol} + 13.9 \text{ kJ/mol} = -21.1 \text{ kJ/mol}\)
23. \(\Delta G\) is negative, so the reaction proceeds spontaneously.
25. \(\Delta G = -RT \ln K_{\text{eq}} = -(8.314 \text{ J/K} \cdot \text{mol})(298 \text{ K})(\ln 1.1 \times 10^{-10}) = -(2477)(-22.9) = 56723 \text{ J/mol} = 56.7 \text{ kJ/mol}\)
26. convert \(\Delta G : 28.9 \text{ kJ/mol} = 28900 \text{ J/mol}\)
27. \(\ln K_{\text{eq}} = \frac{-\Delta G}{-RT} = \frac{28900 \text{ J/mol}}{(8.314)(298)} = -11.7\)
28. and
29. \(K_{\text{eq}} = 8.7 \times 10^{-6}\)
Chapter 20. Entropy and Free Energy Assessments

20.4 Entropy and Free Energy

Chapter Test

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. All of the following are characteristics of exothermic reactions except
   1. energy is absorbed during the reaction
   2. reaction is energetically favorable
   3. products have lower quantity of energy than reactants
   4. products are more stable than reactants

2. All of the following are characteristics of endothermic reactions except
   1. products have higher energy as compared to reactants
   2. heat is absorbed during the reaction
   3. products are more stable than reactants
   4. energy is needed to drive the reaction

3. An increase in entropy is seen in one of the following situations
   1. a lake freezing
   2. making a pizza
   3. shelving books
   4. shredding paper

4. If $\Delta H$ is negative and $\Delta S$ is positive, then $\Delta G$ is
   1. always negative
   2. never negative
   3. negative at higher temperatures
   4. negative at lower temperatures

5. $\text{KOH}(s) \rightarrow K^+(aq) + OH^-(aq) + \text{heat}$ is a spontaneous reaction for all of the following reasons except
   1. entropy is increased
   2. reaction is endothermic
   3. reaction is exothermic
   4. reaction involves change of state

6. At the transition from liquid water to ice
   1. $\Delta G$ for the process is positive
   2. $\Delta G$ for the process is negative
   3. $\Delta H$ for the process is zero
   4. $\Delta S$ for the process is positive

7. When $K_{eq}$ is greater than one, all of the following are true except
   1. $\Delta G$ is negative
   2. $\ln(K_{eq})$ is negative
3. \( \ln(K_{eq}) \) is positive
4. products are favored

8. As the vapor state changes into the liquid state
   1. \( \Delta S \) becomes positive
   2. \( \Delta S \) becomes negative
   3. \( \Delta S = 0 \)
   4. \( \Delta H_{fus} \) changes

9. When the \( H\Delta \) term is numerically larger than the \( T\Delta S \) term, the \( \Delta G \) value will be
   1. positive if \( \Delta H \) is negative
   2. negative if \( \Delta H \) is positive
   3. negative if \( \Delta H \) is negative
   4. variable value

10. The condensation of \( CO_2 \) involves all of the following except
    1. decrease in temperature
    2. change of state
    3. decrease in entropy
    4. increase in volume

True/False:

11. _____ The standard entropy of a substance is given by \( \Delta S^\circ \).
12. _____ Endothermic reactions do not occur spontaneously.
13. _____ The entropy change for the sublimation of \( I_2 \) is positive.
14. _____ Standard conditions for measuring entropy and enthalpy values are 700 mm pressure and 273 K temperature.
15. _____ \( \Delta G \) is negative for a spontaneous reaction.
16. _____ An endothermic reaction has a negative \( \Delta H \).
17. _____ The quicklime process produces measurable amounts of \( CO_2 \) at temperatures above 700°C.
18. _____ When \( K_{eq} = 1, \Delta G = 0 \)
19. _____ \( \Delta G \) can be used to calculate \( K_{sp} \).
20. _____ \( \Delta S_{vap} \) represents the entropy change during the melting process.

Fill in the Blank:

21. Define the following terms:
    1. entropy
    2. spontaneous reaction
    3. nonspontaneous reaction
    4. free energy

22. What is the standard molar entropy change of the following reaction?
23. \( 4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g) \)
24. Given:

\[
\begin{align*}
S_{NH_3}^\circ &= 193 \ J/K\cdot mol \\
S_{O_2}^\circ &= 205 \ J/K\cdot mol \\
S_{NO}^\circ &= 211 \ J/K\cdot mol \\
S_{H_2O}^\circ &= 189 \ J/K\cdot mol
\end{align*}
\]
26. In each of the following cases, indicate whether the process results in an increase or decrease in the entropy of the system and explain with a brief comment.

1. \(C_{10}H_8(l) + 12O_2(g) \rightarrow 10CO_2(g) + 4H_2O(g)\)
2. \(8Fe(s) + S_8(s) \rightarrow 8FeS(s)\)
3. \(Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(s) + H_2(g)\)
4. \(Pb^+(aq) + 2NO_3^-(aq) + 2K^+(aq) + 2I^-(aq) \rightarrow PbI_2(s) + 2K^+(aq) + 2NO_3^-(aq)\)
5. \(2H_2O(l) \rightarrow 2H_2(g) + O_2(g)\)

27. Compute \(G^\circ\) for the hydrolysis of urea, \(CO(NH_2)_2\) at 298.15 K.

28. \(CO(NH_2)_2(aq) + H_2O(l) \rightarrow CO_2(g) + 2NH_3(g)\)

29. \(\Delta H^\circ = 119.2\ kJ/mol\)

30. \(\Delta S^\circ = 354.8\ J/K\cdot mol\)

31. Calculate \(\Delta G^\circ\) for the reaction at 298.15 K

32. \(NH_4NO_3(s) + H_2O(l) \rightarrow NH_4^+(aq) + NO_3^-(aq)\)

33. \(\Delta H^\circ = 28.05\ kJ/mol\)

34. \(\Delta S^\circ = 108.7\ J/K\cdot mol\)

Answer Key

1. a
2. c
3. d
4. a
5. b
6. c
7. b
8. b
9. c
10. d
11. true
12. false
13. true
14. false
15. true
16. false
17. true
18. false
19. true
20. false

1. Entropy is a measure of the degree of randomness or disorder of a system.
2. A spontaneous reaction is a reaction that favors the formation of products at the conditions under which the reaction is occurring.
3. A nonspontaneous reaction is a reaction that does not favor the formation of products at the given set of conditions.
4. Free energy is energy that is available to do work
21. \[ \Delta S_{\text{reaction}} = (4(211 \text{ J/K} \cdot \text{K}) + 6(189 \text{ J/K} \cdot \text{mol})) - (4(193 \text{ J/K} \cdot \text{mol}) + 5(205 \text{ J/K} \cdot \text{mol})) \]
\[ \Delta S_{\text{reaction}} = (844 \text{ J/K} \cdot \text{K} + 1134 \text{ J/K} \cdot \text{mol}) - (772 \text{ J/K} \cdot \text{mol} + 1025 \text{ J/K} \cdot \text{mol}) \]
\[ \Delta S_{\text{reaction}} = 1978 \text{ J/K} \cdot \text{mol} - 1797 \text{ J/K} \cdot \text{mol} \]
\[ \Delta S_{\text{reaction}} = 181 \text{ J/K} \cdot \text{mol} \]

1. increase, more particles in product and change from liquid to gas
2. decrease, lower number of particles in product
3. increase, formation of gas
4. decrease, formation of precipitate
5. increase, more particles in products and formation of gas

22. \[ \Delta S^\circ = 354.8 \text{ J/K} \cdot \text{mol} = 0.3548 \text{ kJ/K} \cdot \text{mol} \]

23. \[ \Delta G^\circ = -119.2 \text{ kJ} - (298.15 \text{ K})(0.3548 \text{ kJ/K}) \]
\[ = +119.2 \text{ kJ} - 105.8 \text{ kJ} \]
\[ = +13.4 \text{ kJ} \]

24. \[ \Delta S^\circ = 108.7 \text{ J/K} \cdot \text{mol} = 0.1087 \text{ kJ/K} \cdot \text{mol} \]

25. \[ \Delta G^\circ = \Delta H^\circ - T \Delta S^\circ \]
\[ = 28.05 \text{ kJ} - (298.15 \text{ K})(0.1087 \text{ kJ/K} \cdot \text{mol}) \]
\[ = 28.050 - 32.410 \]
\[ = -4.360 \text{ kJ/mol} \]
Chapter 21
Acids and Bases Assessments

Chapter Outline

21.1 Acid-Base Definitions
21.2 The pH Concept
21.3 Acid and Base Strength
21.4 Acid-Base Neutralization
21.5 Salt Solutions
21.6 Acids and Bases
21.1 Acid-Base Definitions

Lesson Quiz

Multiple Choice:

1. All of the following are bases except
   1. drain cleaner
   2. ammonia
   3. vinegar
   4. sodium hydrogen carbonate

2. One of the following is a property of acids
   1. turn litmus red
   2. bitter taste
   3. do not react with metals
   4. turn phenolphthalein pink

3. One of the following is a property of bases
   1. sour taste
   2. aqueous solutions are electrolytes
   3. turn litmus blue
   4. react with some metals to produce H₂

4. H₃O⁺ is formally known as the
   1. hydrogen ion
   2. dihydrogen monoxide
   3. hydronium ion
   4. protonated water

5. Phosphoric acid is a ______ acid
   1. monoprotic
   2. diprotic
   3. multiprotic
   4. polyprotic

6. Alkali metals react with water to form
   1. base + hydroxide compound
   2. base + CO₂
   3. base + oxygen
   4. base + hydrogen

7. Ammonia is classified as a base because of all the following except
   1. turns litmus blue
   2. forms hydroxide ion
3. reacts with acids
4. turns phenolphthalein pink

8. In the reaction between water and ammonia, water functions as a
   1. Brønsted-Lowry acid
   2. Brønsted-Lowry base
   3. Lewis acid
   4. Arrhenius base

9. A Lewis acid
   1. accepts an electron pair
   2. donates an electron pair
   3. donates a proton
   4. donates a hydroxide ion

10. A proton is a
    1. hydroxide ion
    2. hydrogen ion
    3. hydronium ion
    4. hydrated ion

True/False:

11. _____ Oranges contain citric acid
12. _____ Lithium carbonate is a commercially available antacid
13. _____ All bases are strong electrolytes
14. _____ Acids react with active metals to produce hydrogen gas.
15. _____ CH₃OH is a strong base.
16. _____ Water can function as both a Brønsted-Lowry acid and a Brønsted-Lowry base.
17. _____ In a Lewis acid-base reaction, a covalent bond is formed.
18. _____ Acids are often called proton donors.
19. _____ The sulfate ion is the conjugate acid of the hydrogen sulfate ion.
20. _____ Group 2 metal hydroxides completely dissolve in water.

Fill in the blanks:

21. A Brønsted-Lowry _____ is a molecule or ion that donates a ______ ion in a reaction.
22. Water can also act as a base in a ______ acid-base reaction, as long as it reacts with a substance that is a ______ proton donor.
23. An ______ substance is one that is capable of acting as either an _____ or a base by donating or ______-hydrogen ions.
24. A conjugate ______ is the particle produced when a base ______ a proton.
25. Name the following compounds:
   1. NaOH
   2. HClO
   3. H₃PO₄
   4. Mg(OH)₂

26. Write the formulas for the following compounds:
   1. calcium hydroxide
   2. sulfuric acid
   3. potassium hydroxide
21.1. Acid-Base Definitions

4. acetic acid

27. Identify the Lewis acid and the Lewis base in each of the following reactions:
   1. \( \text{NH}_3 + BF_3 \rightarrow H_3N - BF_3 \)
   2. \( CH_3 - O^- + CH_3Cl \rightarrow CH_3 - O - CH_3 + Cl^- \)
   3. \( (CH_3)_3O^+ + H - O - H \rightarrow CH_3 - O - CH_3 + CH_3OH^2+ \)
   4. \( Cu^{2+} + 4NH_3 \rightarrow Cu(NH_3)_4^{+2} \)

---

**Answer Key**

1. c
2. a
3. c
4. c
5. d
6. d
7. b
8. a
9. a
10. b
11. true
12. false
13. false
14. true
15. false
16. true
17. true
18. true
19. false
20. false
21. acid, hydrogen
22. Brønsted-Lowry, better
23. amphoteric, acid, accepting
24. acid, accepts
   1. sodium hydroxide
   2. hypochlorous acid
   3. phosphoric acid
   4. magnesium hydroxide
   1. Ca(OH)\(_2\)
   2. H\(_2\)SO\(_4\)
   3. KOH
   4. CH\(_3\)COOH

---

**Table 21.1:**

<table>
<thead>
<tr>
<th>Problem</th>
<th>Lewis Acid</th>
<th>Lewis Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>a.</td>
<td>NH(_3)</td>
<td>BF(_3)</td>
</tr>
<tr>
<td>b.</td>
<td>CH(_3)Cl</td>
<td>CH(_3)-O(^-)</td>
</tr>
<tr>
<td>Problem</td>
<td>Lewis Acid</td>
<td>Lewis Base</td>
</tr>
<tr>
<td>---------</td>
<td>---------------</td>
<td>------------</td>
</tr>
<tr>
<td>c.</td>
<td>(CH₃)₃O⁺</td>
<td>H–O–H</td>
</tr>
<tr>
<td>d.</td>
<td>Cu⁺²</td>
<td>NH₃</td>
</tr>
</tbody>
</table>
21.2 The pH Concept

Lesson Quiz

Name___________________ Class________________ Date________

Multiple Choice:

1. $\text{H}_3\text{O}^+$ is the structure of a
   1. hydrated proton
   2. hydronium ion
   3. protonated water
   4. hydrogen ion

2. $1 \times 10^{-14}$ is the equilibrium constant for
   1. ionization of weak base
   2. ionization of weak acid
   3. ionization of water
   4. ionization of acetic acid

3. The pH for a strong acid is
   1. greater than 7
   2. greater than 9
   3. 7
   4. less than 7

4. The most acid solution in the following list is
   1. tomato juice
   2. lemon juice
   3. soda pop
   4. coffee

5. The most basic solution in the following list is
   1. household bleach
   2. eggs
   3. soap
   4. ammonia solution

6. A solution with a pH of 9 has a hydrogen ion concentration of
   1. $1 \times 10^{-8} \text{ } M$
   2. $9 \times 10^{-1} \text{ } M$
   3. $1 \times 10^{-9} \text{ } M$
   4. $9 \times 10^{-9} \text{ } M$

7. The pH of a solution is 4, so the pOH is
   1. 11
   2. 9
3. 13
4. 10

8. The pH of a solution can be calculated by
   1. \( pH = \log[H^+] \)
   2. \( pH = \ln[H^+] \)
   3. \( pH = -\log[H^+] \)
   4. \( pH = -\ln[H^+] \)

9. Soap has an approximate pH of 11. Its \([H^+]\) is approximately
   1. \( 1 \times 10^{11} M \)
   2. \( 1 \times 10^{-3} M \)
   3. \( 1 \times 10^{-7} M \)
   4. \( 1 \times 10^{-11} M \)

10. All of the following have pH values greater than 8 except
    1. blood
    2. detergent
    3. soap
    4. milk of magnesia

True/False:

11. _____ Water is a weak electrolyte.
12. _____ The \( K_w \) is affected by temperature.
13. _____ When NaOH is added to water, the equilibrium shifts to the \( H_2O \) (l) side.
14. _____ A 0.3 M solution of HCl has a negative pH.
15. _____ The letters pH stand for power of the hydrogen ion.
16. _____ The pH of vinegar is greater than 3.
17. _____ Soap has an approximate \([H^+]\) of \( 10^{-11} \).
18. _____ The pOH value = 14 + pH.
19. _____ Hydrogen ions exist free in water solution.
20. _____ The equilibrium position for the self-ionization of water strongly favors the reactant water molecule.

Fill in the blanks:

21. The ______ of water is the process in which water ______ to ______ ions and hydroxide ions.
22. The ______ of water (Kw) is the mathematical ______ of the concentrations of hydrogen ions and hydroxide ions.
23. An ______ solution is a solution in which the concentration of hydrogen ions is greater than the concentration of ______ions.
24. The pH scale is generally presented as running from 0 to 14, though it is possible to have a pH of less than ___ or greater than ___
25. State whether the following solutions are acidic or basic:
   1. \([H^+] = 3.7 \times 10^{-4} M\)
   2. \([H^+] = 1.9 \times 10^{-9} M\)
   3. \([OH^-] = 5.8 \times 10^{-8} M\)
   4. \([OH^-] = 2.5 \times 10^{-5} M\)
26. Calculate \([H^+]\) and \([OH^-]\) for the following solutions which ionize completely in water:
   1. 0.07 M HBr
   2. \(1.4 \times 10^{-4} M \ HCl\)
21.2. The pH Concept

3. 0.028 M NaOH
4. $3.6 \times 10^{-5}$ M LiOH

27. Calculate the pH and pOH of the following solutions:

1. $[H^+] = 4.6 \times 10^{-4}$ M
2. $[OH^-] = 0.0063$ M

---

**Answer Key**

1. b
2. c
3. d
4. b
5. a
6. c
7. d
8. c
9. d
10. a
11. true
12. true
13. true
14. false
15. true
16. false
17. true
18. false
19. false
20. true
21. self-ionization, ionizes, hydronium
22. ion-product, product
23. acidic, hydroxide
24. 0, 14
   1. acidic
   2. basic
   3. acidic
   4. basic

   a. $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$  
      $0.07$ M = $[H^+]$
   b. $[OH^-] = \frac{K_w}{[H^+]} = \frac{1 \times 10^{-14}}{7 \times 10^{-2}} = 1.43 \times 10^{-13}$
   c. $[H^+] = 1.4 \times 10^{-4}$ M and $[OH^-] = \frac{K_w}{[H^+]} = \frac{1 \times 10^{-14}}{1.4 \times 10^{-4}} = 7.14 \times 10^{-11}$ M
   d. $[OH^-] = 2.8 \times 10^{-2}$ M and $[H^+] = \frac{K_w}{[OH^-]} = \frac{1 \times 10^{-14}}{2.8 \times 10^{-2}} = 3.57 \times 10^{-13}$ M
   e. $[OH^-] = 3.6 \times 10^{-5}$ M and $[H^+] = \frac{K_w}{[OH^-]} = \frac{1 \times 10^{-14}}{3.6 \times 10^{-5}} = 2.78 \times 10^{-10}$ M

   a. $pH = -\log[H^+] = -\log(4.6 \times 10^{-4}$ M) = 3.33
   b. $pOH = 14 - pH = 14 - 3.33 = 10.67$
   c. $pOH = -\log[OH^-] = -\log(0.0063) = 2.20$
   d. $pH = 14 - pOH = 14 - 2.20 = 11.8$
21.3 Acid and Base Strength

Lesson Quiz

Name___________________ Class_________________ Date________

Multiple Choice:

1. All of the following are weak acids except
   1. phosphoric acid
   2. carbonic acid
   3. hydrocyanic acid
   4. sulfuric acid

2. Strong acids ionize to form
   1. protons and anions
   2. protons and a conjugate base
   3. protons and a conjugate acid
   4. a conjugate acid and a conjugate base

3. Based on their first acid ionization constant, the strongest acid is
   1. oxalic acid
   2. benzoic acid
   3. acetic acid
   4. nitrous acid

4. Based on the base ionization constants, the weakest base is
   1. methylamine
   2. urea
   3. ammonia
   4. pyridine

5. The base ionization represents all of the following except
   1. the extent of ionization
   2. the strength of the base
   3. an indication of undissociated base
   4. the number of anions present

6. The weakest conjugate base listed below is
   1. CN⁻
   2. HSO₄⁻
   3. CH₃COO⁻
   4. HCO₃⁻

7. Only one of the following is a weak acid
   1. H₂CO₃
   2. HCl
3. HBr
4. HNO₃

8. The base ionization constant contains all of the following except
   1. unionized base
   2. base cation
   3. water
   4. hydroxide anion

9. The extent of ionization for a weak acid is usually less than
   1. 90%
   2. 50%
   3. 10%
   4. 0%

10. The base ionization constant expression for ammonia is $K_b =$
    1. $[NH_4^+][OH^-] / [NH_3]$
    2. $[NH_3] / [NH_4^+][OH^-]$
    3. $[OH^-] / [NH_4^+][NH_3]$
    4. $[NH_4^+] / [OH^-][NH_3]$

**True/False:**

11. _____ Some strong acids are found in foods and beverages.
12. _____ Hydrofluoric acid is a weak acid.
13. _____ HCO₃⁻ is the conjugate base of carbonic acid.
14. _____ The third ionization constant for phosphoric acid is larger than the first.
15. _____ Potassium hydroxide is completely soluble in water.
16. _____ The $K_b$ for a weak base has $[OH^-]$ in the denominator.
17. _____ Methylamine is a weaker base than pyridine.
18. _____ Calculations of $K_a$ omit the $[H^+]$ that is formed by the ionization of water.
19. _____ The conjugate base of a weak acid is a weak base.
20. _____ When dissolved in water, ammonia donates a proton.

**Fill in the blanks:**

21. A _____ acid is an acid which is completely ________in ________ solution.
22. The ________ value of $K_a$ is a reflection of the ________of the acid.
23. Each successive ________constant for a polyprotic acid is always ________than the previous one.
24. Write an acid ionization reaction for each of the following acids:
   1. HCOOH
   2. H₂CrO₄
   3. HClO
25. Write a base ionization reaction for each of the following bases:
   1. CH₃CH₂NH₂
   2. C₅H₅N
   3. C₆H₅NH₂
26. Calculate the pH of the following solutions
   1. 0.062 M acetic acid ($K_a = 1.8 \times 10^{-5}$)
   2. $1.55 \times 10^{-4}$ M $HClO_4$
   3. 0.8 M trimethylamine ($K_b = 7.4 \times 10^{-5}$)
27. Complete the following table:
### Table 21.2:

<table>
<thead>
<tr>
<th></th>
<th>pH</th>
<th>([H^+])</th>
<th>pOH</th>
<th>([\text{OH}^-])</th>
</tr>
</thead>
<tbody>
<tr>
<td>a.</td>
<td>5.4 \times 10^{-4}</td>
<td></td>
<td></td>
<td>0</td>
</tr>
<tr>
<td>b.</td>
<td>10.75</td>
<td></td>
<td></td>
<td>7.8 \times 10^{-10}</td>
</tr>
<tr>
<td>c.</td>
<td>7.8 \times 10^{-10}</td>
<td></td>
<td></td>
<td>8.3</td>
</tr>
<tr>
<td>d.</td>
<td>5.00</td>
<td></td>
<td></td>
<td>7.0</td>
</tr>
</tbody>
</table>

### Answer Key

1. d  
2. a  
3. a  
4. b  
5. d  
6. b  
7. a  
8. c  
9. c  
10. a  
11. false  
12. true  
13. true  
14. false  
15. true  
16. false  
17. false  
18. true  
19. true  
20. false  
21. strong, ionized, aqueous  
22. numerical, strength  
23. ionization, smaller  

   1. \(\text{HCOOH} \rightleftharpoons H^+ + COO^-\)  
   2. \(\text{H}_2\text{CrO}_4 \rightleftharpoons H^+ + \text{HCrO}_4^-\) (first ionization)  
   3. \(\text{HClO} \rightleftharpoons H^+ + \text{ClO}^-\)

   1. \(\text{CH}_3\text{CH}_2\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{CH}_2\text{NH}_3^+ + \text{OH}^-\)  
   2. \(\text{C}_5\text{H}_5\text{N} + \text{H}_2\text{O} \rightleftharpoons \text{C}_5\text{H}_5\text{NH}^+ + \text{OH}^-\)  
   3. \(\text{C}_6\text{H}_5\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{C}_6\text{H}_5\text{NH}_3^+ + \text{OH}^-\)

24. a. \(\text{HAc} \rightleftharpoons H^+ + \text{Ac}^-\)

### Table 21.3:

<table>
<thead>
<tr>
<th></th>
<th>[HAc]</th>
<th>[H(^+)]</th>
<th>[Ac(^-)]</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>0.062</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>(-x)</td>
<td>x</td>
<td>(x)</td>
</tr>
<tr>
<td>E</td>
<td>0.062(-x)</td>
<td>x</td>
<td>(x)</td>
</tr>
</tbody>
</table>
\[1.8 \times 10^{-5} = \frac{(x)(x)}{0.062^2} \approx \frac{x^2}{0.062} \text{ then } x^2 = 1.16 \times 10^{-6} \text{ and } x = 1.08 \times 10^{-3}\]

\[pH = -\log(1.08 \times 10^{-3}) = 2.97\]

b. HClO₄ is a strong acid so \([H^+] = 1.55 \times 10^{-4}\) \(pH = -\log[H^+] = 3.81\)

c. Let \(TM = \text{trimethylamine}, \) then \(TM + H_2O \rightleftharpoons TMH^+ + OH^-\)

### Table 21.4:

<table>
<thead>
<tr>
<th></th>
<th>[TM]</th>
<th>[TMH⁺]</th>
<th>[OH⁻]</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>0.8 M</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>C</td>
<td>−x</td>
<td>x</td>
<td>x</td>
</tr>
<tr>
<td>E</td>
<td>0.8 − x</td>
<td>x</td>
<td>x</td>
</tr>
</tbody>
</table>

\[7.4 \times 10^{-5} = \frac{(x)(x)}{0.8 − x} \approx \frac{x^2}{0.8} \text{ then } x^2 = 5.92 \times 10^{-5} \text{ and } x = 7.7 \times 10^{-3} \text{ and } pOH = 2.11\]

\[pH = 14 − 2.11 = 11.89\]

### Table 21.5:

<table>
<thead>
<tr>
<th></th>
<th>pH</th>
<th>([H^+])</th>
<th>pOH</th>
<th>([OH^-])</th>
</tr>
</thead>
<tbody>
<tr>
<td>a.</td>
<td>3.27</td>
<td>5.4 \times 10^{-4}</td>
<td>10.73</td>
<td>1.9 \times 10^{-11}</td>
</tr>
<tr>
<td>b.</td>
<td>4.89</td>
<td>1.3 \times 10^{-5}</td>
<td>9.11</td>
<td>7.8 \times 10^{-10}</td>
</tr>
<tr>
<td>c.</td>
<td>10.75</td>
<td>1.8 \times 10^{-11}</td>
<td>3.25</td>
<td>5.6 \times 10^{4}</td>
</tr>
<tr>
<td>d.</td>
<td>9.00</td>
<td>1.0 \times 10^{-9}</td>
<td>5.00</td>
<td>1.0 \times 10^{-5}</td>
</tr>
</tbody>
</table>
21.4 Acid-Base Neutralization

Lesson Quiz

Name___________________ Class________________ Date________

Multiple Choice:

1. In the reaction $H^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl^- + H_2O$, the spectator ions are
   1. $H^+ + Cl^-$
   2. $Na^+ + OH^-$
   3. $Na^+ + Cl^-$
   4. $H^+ + OH^-$

2. The neutralization of a weak base with a strong acid produces a solution whose pH is
   1. neutral
   2. slightly acidic
   3. slightly basic
   4. strongly basic

3. The neutralization of a strong base with a strong acid produces a solution whose pH is
   1. neutral
   2. slightly acidic
   3. slightly basic
   4. strongly basic

4. In a titration of sulfuric acid with KOH, the ratio of acid to base for calculations is
   1. 1 mol sulfuric acid/1 mol NaOH
   2. 2 mol sulfuric acid/1 mol NaOH
   3. 1 mol sulfuric acid/2 mol NaOH
   4. 1 mol sulfuric acid/3 mol NaOH

5. After the neutralization point is reached in a titration, the pH
   1. changes very slowly
   2. changes rapidly
   3. shows no change
   4. fluctuates depending on the acid

6. If the equivalence point is greater than pH seven, the titration is between
   1. a strong acid and a strong base
   2. a weak acid and a strong base
   3. a strong base and a weak acid
   4. a weak base and a weak acid

7. Thymol Blue exhibits a color change between the approximate pH values of
   1. 2-3
   2. 4-6
3. 1-3
4. 10-12

8. The best indicator for a titration with an equivalence point of 10 would be
   1. thymol blue
   2. thymolphthalein
   3. phenolphthalein
   4. alizarine yellow R

9. Red cabbage juice has a _____ color in neutral solutions
   1. red
   2. green
   3. purple
   4. orange

10. Antacids neutralize stomach HCl by using a material such as
    1. NaOH
    2. HNO₃
    3. NH₃
    4. Ca(OH)₂

True/False:

11. _____ Net ionic equations of weak bases include the weak base as a molecule.
12. _____ In neutralization reactions between weak acids and strong bases, both the acid and the base are completely ionized.
13. _____ The net ionic equation for the neutralization of a weak acid by a strong base is $HA + OH^- \rightarrow A^- + H₂O$.
14. _____ In a titration, the material from the buret is added rapidly to achieve equilibrium.
15. _____ A titration curve plots volume of titrant against pH.
16. _____ Titration curves for weak acid-strong base titrations have a different equivalence point when compared to strong acid-strong base titrations.
17. _____ The acid that aids in the digestion of food is phosphoric acid.
18. _____ The products of a typical acid-base neutralization are a salt plus water.
19. _____ Phosphoric acid requires two moles NaOH/mole acid for neutralization.
20. _____ In a titration between HCl and ammonia, the ammonia functions as a proton acceptor.

Fill in the blanks:

21. A ________reaction is a reaction in which an ______and a base react in an ________ solution to produce a salt and______.
22. A salt is an ionic compound composed of a ______ from a base and an anion from an _____.
23. Reactions where at least one of the components is _______do not generally result in a ________ solution.
24. The ________ point is the point in a ________reaction where the number of ______of hydrogen ions is equal to the number of moles of ______ions.
25. Which indicator would be the best to use for a titration that has an equivalence point at pH 10.0?
26. In a titration, 22.5 mL of 0.4 M NaOH were required to neutralize 15.3 mL of HCl. What is the concentration of the HCl?
27. It takes 35.7 mL of 0.05 M NaOH to neutralize 21.6 mL of phosphoric acid. What is the concentration of the phosphoric acid?
28. What would the color of phenol red be in a solution at pH 3? pH 7.8?
29. Sulfuric acid contains two protons that dissociate under different conditions. The first dissociation (to form $H^+ + HSO_4^-$) is listed as “very large”, equivalent to that of HCl. The second dissociation (to form $H^+ + SO_4^{2-}$) has a $K_a$ of $1.9 \times 10^{-2}$. What would you expect a titration curve to look like for sulfuric acid?
21.4. Acid-Base Neutralization

Answer Key

1. c
2. b
3. a
4. c
5. a
6. c
7. c
8. b
9. c
10. d
11. true
12. false
13. true
14. false
15. true
16. true
17. false
18. true
19. false
20. true
21. neutralization, acid, aqueous, water
22. cation, acid
23. weak, neutral
24. equivalence, neutralization, moles, hydroxide
25. thymolphthalein
26. $M_A \times V_A = M_B \times V_B$ so $M_a = \frac{(M_b)(V_b)}{V_a} = \frac{(0.4 \, M)(22.5 \, mL)}{15.3 \, mL} = 0.59 \, M$
27. moles NaOH = 37.7 mL $\times \frac{0.05 \, moles}{L} \times \frac{L}{1000 \, mL} = 1.89 \times 10^{-3} \, moles$
28. phosphoric acid has 3 hydrogens/molecule, so
29. $(1.89 \times 10^{-3} \, moles \, NaOH) \times \frac{1 \, mole \, H_3PO_4}{3 \, moles \, NaOH} = 6.3 \times 10^{-4} \, moles \, H_3PO_4$
30. phosphoric acid concentration = $6.3 \times 10^{-4} \, moles \div 21.6 \, mL \times 1000 \, mL/L = 0.029 \, M$ phosphoric acid
31. 
32. At pH 3, the indicator would be colorless. At pH 7.8 it would be orange-red.
33. The curve would probably resemble that of HCl since the first $K_a$ is so large.
Lesson Quiz

Name___________________ Class______________ Date________

Multiple Choice:

1. When CsNO₃ is dissolved in water, the resulting solution will be
   1. strongly basic
   2. slightly basic
   3. neutral
   4. slightly acidic

2. A solution of RbF in water will have a pH that is
   1. strongly basic
   2. slightly basic
   3. neutral
   4. slightly acidic

3. A solution of a salt from a weak acid and a strong base will be
   1. strongly basic
   2. slightly basic
   3. neutral
   4. slightly acidic

4. A solution containing acetic acid \(K_a = 1.8 \times 10^{-5}\) and urea \(K_b = 1.5 \times 10^{-14}\) will be
   1. acidic
   2. basic
   3. too difficult to analyze
   4. neutral

5. The acetate ion can react with water by
   1. being a proton donor
   2. being a proton acceptor
   3. forming a hydroxide salt
   4. neutralizing the hydroxide ion

6. The nitrogen-containing base pyridine can react with water by serving as all of the following except
   1. proton donor
   2. proton acceptor
   3. Brønsted-Lowry base
   4. Lewis acid

7. The CN⁻ anion will produce an aqueous salt solution that is
   1. slightly basic
   2. slightly acid
3. neutral
4. unable to predict

8. Fluoride in water acts as all of the following except
   1. Arrhenius base
   2. Brønsted-Lowry acid
   3. Brønsted-Lowry base
   4. Lewis base

9. A buffer composed of acetic acid and sodium acetate is an example of
   1. weak acid and salt of its conjugate base
   2. weak base and salt of its conjugate acid
   3. strong acid and salt of its conjugate base
   4. strong base and salt of its conjugate acid

10. The ammonia/ammonium ion buffer is an example of
   1. weak acid and salt of its conjugate base
   2. weak base and salt of its conjugate acid
   3. strong acid and salt of its conjugate base
   4. strong base and salt of its conjugate acid

True/False:

11. _____ The ammonium ion reacts with water to lower the pH of the solution.
12. _____ Anions of weak acids contribute to proton formation.
13. _____ The conjugate base of HCN will produce OH\(^-\) ions.
14. _____ NO\(_2^-\) is the conjugate base of nitric acid.
15. _____ Sodium acetate acts as a base in water.
16. _____ A solution becomes basic as a result of hydronium ion production.
17. _____ The bromide ion will attract protons from water.
18. _____ The ammonia/ammonium ion buffer is important for maintaining blood pH.
19. _____ The acetic acid/acetate buffer functions best at pH values above 7.
20. _____ A salt formed from a weak base and a strong acid will be acidic.

Fill in the blanks:

21. Predict whether solutions made from the following salts will be acidic, basic, or neutral:
   1. KOH
   2. NaHCO\(_3\)
   3. NH\(_4\)Cl

22. A _____ is an _____ compound that is formed when an acid and a base _____ each other.
23. Salt _____ is a reaction in which one of the _____ from a salt reacts with_______, forming either an acidic
    or basic solution.
24. Salts that are derived from the neutralization of a _____ acid (HF) by a _____ base (NaOH) will always
    produce salt solutions that are _______.
25. The buffer _____ is the amount of acid or base that can be added to a _______ solution before a large change
    in ____ occurs.
26. Write balanced equations for the following ions in water solution:
   1. NO\(_2^-\)
   2. C\(_6\)H\(_5\)COO\(^-\)
   3. H\(_2\)NCONH\(_2\)
27. What is the pH of a 0.01 M solution of benzoic acid? \( K_a = 6.5 \times 10^{-5} \). Benzoic acid dissociates according to the following equation:

\[ C_6H_5COOH \rightarrow H^+ + C_6H_5COO^- \]

28. 

29. 

30. What would the pOH of this solution be?

---

### Answer Key

1. c  
2. b  
3. b  
4. c  
5. b  
6. a  
7. a  
8. b  
9. a  
10. b  
11. false  
12. false  
13. true  
14. false  
15. true  
16. false  
17. false  
18. false  
19. false  
20. true

1. basic  
2. basic  
3. acidic

21. salt, ionic, neutralize  
22. hydrolysis, ions, water  
23. weak, strong, basic  
24. capacity, buffered, pH

1. \( \text{NO}_2 + H_2O \rightleftharpoons \text{HNO}_2 + OH^- \)  
2. \( C_6H_5COO^- + H_2O \rightleftharpoons C_6H_5COOH + OH^- \)  
3. \( H_2NCONH_2 + H_2O \rightleftharpoons H_2NCONH_3 + OH^- \)  
4. \( \text{CH}_3\text{NH}_3^+ + H_2O \rightleftharpoons \text{H}_3\text{O}^+ + \text{CH}_3\text{NH}_2 \)

25. \( C_6H_5COOH \rightarrow H^+ + C_6H_5COO^- \)

26. Let \( \text{BA} = \text{benzoic acid} \)  
28. and \( B = \text{benzoate anion} \)

30. then \( K_a = \frac{[H^+][B^-]}{[BA]} \)
32. we know that
33. \([H^+] = [B^-]\)
34. so numerator can be
35. \((x)(x)\)
36. or
37. \(x^2\)
38. equilibrium concentration of BA is
39. \(0.01 \, M - x \approx 0.01 \, M\)
40. so
41. \(x^2 = (6.5 \times 10^{-5})(1 \times 10^{-2}) = 65 \times 10^{-8}\)
42. and
43. \(x = 8.06 \times 10^{-4}\)
44. \(pH = -\log 8.06 \times 10^{-4} = 3.06\)
45. \(pH = 3.06\) then \(pOH = 14 - pH = 10.94\)
21.6 Acids and Bases

Chapter Test

Name___________________ Class______________ Date________

Multiple Choice:

1. Bases have all of the following properties except
   1. turn litmus red
   2. bitter taste
   3. do not react with metals
   4. turn phenolphthalein pink

2. A Lewis base
   1. accepts an electron pair
   2. donates an electron pair
   3. donates a proton
   4. donates a hydroxide ion

3. All of the following have basic pH values except
   1. soap
   2. lemon juice
   3. detergent
   4. eggs

4. The pOH of a solution is 6.9 so the pH is
   1. 6.8
   2. 7.6
   3. 7.1
   4. 7.4

5. All of the following are strong acids except
   1. HCl
   2. HNO₂
   3. HNO₃
   4. HBr

6. The strongest conjugate base listed below is
   1. CN⁻
   2. HSO₄⁻
   3. CH₃COO⁻
   4. HCO₃⁻

7. The neutralization of a weak acid with a strong base produces a solution whose pH is
   1. neutral
   2. slightly acidic
3. slightly basic
4. strongly basic

8. In a titration of hydrobromic acid with KOH, the ratio of acid to base for calculations is
   1. 1 mol hydrobromic acid/1 mol NaOH
   2. 2 mol hydrobromic acid/1 mol NaOH
   3. 1 mol hydrobromic acid/2 mol NaOH
   4. 1 mol hydrobromic acid/3 mol NaOH

9. NH$_3$ in water acts as all of the following except
   1. Arrhenius acid
   2. Lewis acid
   3. Arrhenius base
   4. Brønsted-Lowry acid

10. The carbonic acid/hydrogen carbonate ion buffer is an example of
   1. weak acid and salt of its conjugate base
   2. weak base and salt of its conjugate acid
   3. strong acid and salt of its conjugate base
   4. strong base and salt of its conjugate acid

True/False:

11. _____ Bases react with active metals to produce hydrogen gas.
12. _____ Ca(OH)$_2$ dissolves completely in water.
13. _____ HBr is a strong electrolyte.
14. _____ The pH of soda pop is approximately 7.
15. _____ CO$_3^{2-}$ is the conjugate base of carbonic acid.
16. _____ When dissolved in water, ammonia accepts a proton.
17. _____ Net ionic equations of weak acids include the weak acid as a molecule.
18. _____ Titration curves for strong acid-weak base titrations have a different equivalence point when compared to strong acid-strong base titrations.
19. _____ NO$_2^-$ is the conjugate base of nitrous acid.
20. _____ A salt formed from a strong base and a weak acid will be acidic.

Fill in the blanks:

21. Define the following terms:
   1. polyprotic acid
   2. Brønsted-Lowry acid
   3. amphoteric
   4. salt

22. Classify each of the following as an Arrhenius acid/base (A), a Brønsted-Lowry acid/base (BL), or a Lewis acid/base (L)

<table>
<thead>
<tr>
<th>Compound</th>
<th>Classification</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td></td>
</tr>
<tr>
<td>Mg(OH)$_2$</td>
<td></td>
</tr>
<tr>
<td>H$_2$O</td>
<td></td>
</tr>
<tr>
<td>NH$_3$</td>
<td></td>
</tr>
<tr>
<td>BF$_3$</td>
<td></td>
</tr>
</tbody>
</table>
23. State whether the following solutions are acidic or basic:
   1. \([H^+] = 3.9 \times 10^{-4} M\)
   2. \([OH^-] = 5.7 \times 10^{-6} M\)
   3. \([H^+] = 9.3 \times 10^{-10} M\)
   4. \([OH^-] = 6.8 \times 10^{-8} M\)

24. Calculate \([H^+]\) and \([OH^-]\) for the following solutions which ionize completely in water:
   1. 0.09 M HI
   2. \(2.7 \times 10^{-3} M\) HBr
   3. 0.08 M KOH
   4. \(8.4 \times 10^{-4} M\) CsOH

25. Calculate the pH and pOH of the following solutions:
   1. \([H^+] = 6.3 \times 10^{-5} M\)
   2. \([OH^-] = 0.0037 M\)

26. Write an acid or base ionization reaction for the following compounds:
   1. HCOOH
   2. N(CH\(_3\))\(_3\)
   3. H\(_2\)S
   4. HS\(^-\)

27. Like any equilibrium constant, \(K_w\) varies with temperature. Its value at 37\(^\circ\)C is \(2.4 \times 10^{-14}\). What is the pH of (neutral) pure water at 37\(^\circ\)C?

28. In a titration, 31.6 mL of 0.45 M NaOH were required to neutralize 23.2 mL of HCl. What is the concentration of the HCl?

---

**Answer Key**

1. a
2. b
3. b
4. c
5. b
6. a
7. c
8. a
9. c
10. a
11. false
12. false
13. true
14. false
15. false
16. true
17. true
18. true
19. true
20. false

1. an acid that contains multiple ionizable hydrogens.
2. a molecule or ion that donates a hydrogen ion in a reaction.
3. An amphoteric substance is one that is capable of acting as either an acid or a base by donating or accepting hydrogen ions.
4. an ionic compound composed of a cation from a base and an anion from an acid.

<table>
<thead>
<tr>
<th>Table 21.7:</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Compound</strong></td>
</tr>
<tr>
<td>HCl</td>
</tr>
<tr>
<td>Mg(OH)₂</td>
</tr>
<tr>
<td>H₂O</td>
</tr>
<tr>
<td>NH₃</td>
</tr>
<tr>
<td>BF₃</td>
</tr>
</tbody>
</table>

23. 1. acidic
2. basic
3. basic
4. acidic

a. \([H^+] = 9 \times 10^{-2} \text{ M}\)
b. \([H^+][OH^-] = 1 \times 10^{-14} \text{ so } [OH^-] = \frac{1 \times 10^{-14}}{9 \times 10^{-2}} = 1.11 \times 10^{-13} \text{ M}\)
c. \([H^+] = 2.7 \times 10^{-3} \text{ M so } [OH^-] = \frac{1 \times 10^{-14}}{2.7 \times 10^{-3}} = 3.70 \times 10^{-12} \text{ M}\)
d. \([OH^-] = 0.08 \text{ M so } [H^+] = \frac{1 \times 10^{-14}}{8 \times 10^{-4}} = 1.25 \times 10^{-13} \text{ M}\)
e. \([OH^-] = 8.4 \times 10^{-4} \text{ M so } [H^+] = \frac{1 \times 10^{-14}}{8.4 \times 10^{-4}} = 1.19 \times 10^{-11} \text{ M}\)

1. \(pH = -\log(6.3 \times 10^{-5}) = 4.2 \quad pOH = 14 - pH = 14 - 4.2 = 9.8\)
2. \(pOH = -\log(0.0037) = 2.4 \quad pH = 14 - 2.4 = 11.6\)

1. \(HCOOH \rightleftharpoons H^+ + HCOO^-\)
2. \(N(CH_3)_3 + H_2O \rightleftharpoons ^+HN(CH_3)_3 + OH^-\)
3. \(H_2S \rightleftharpoons H^+ + HS^-\)
4. \(HS^- + H_2O \rightleftharpoons H_2S + OH^-\)

24. \(K_w = [H^+][OH^-]\)
25. since
26. \([H^+] = [OH^-]\)
27. we can write
28. \(2.4 \times 10^{-14} = [H^+]^2\)
29. then
30. \([H^+] = 1.2 \times 10^{-7} \text{ M}\)
31. and
32. \(pH = -\log(1.2 \times 10^{-7}) = 6.9\)
33. \(V_aM_a = V_bM_b \text{ so } M_a = \frac{V_bM_b}{V_a} = (31.6 \text{ mL})(0.45 \text{ M}) \div 23.2 \text{ mL} = 0.61 \text{ M HCl}\)
Chapter Outline

22.1 The Nature of Oxidation and Reduction
22.2 Oxidation Numbers
22.3 Balancing Redox Reactions
22.4 Oxidation-Reduction Reactions
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. One of the following is not an oxidation process
   1. formation of MgO
   2. killing bacteria with H₂O₂
   3. refining iron ore
   4. rusting of iron

2. One of the following is not a reduction process
   1. removal of oxygen
   2. reacting Fe₂O₃ with carbon
   3. heating mercuric oxide
   4. bleaching stains

3. Reduction involves
   1. loss of electrons
   2. shifting electrons from oxygen to another atom
   3. loss of hydrogen
   4. loss of protons

4. When oxygen bonds to fluorine
   1. both atoms form ions
   2. electrons shift from oxygen to fluorine
   3. electrons shift from fluorine to oxygen
   4. electrons are shared equally between the two atoms

5. When Zn is oxidized
   1. electrons are added to the Zn atom
   2. electrons are lost into the solution
   3. gaseous Zn is formed
   4. the electrons transfer from Zn to another atom

6. In water, bonding electrons are more attracted to O than H because
   1. the O molecule has a double bond structure
   2. O has a higher electronegativity than H
   3. H has a higher electronegativity that O
   4. the reactants are completely nonpolar

7. In reactions involving molecular compounds, all of the following are true except
   1. electrons are completely transferred between atoms
   2. electrons shift toward one atom in a covalent bond
3. a partial charge can form on an atom in a covalent bond
4. one atom in a covalent bond has less electron density than the other atom

8. Rust is a complex mixture of
   1. iron halides
   2. iron phosphates
   3. iron hydroxides
   4. iron oxides

9. One metal that does not corrode easily is
   1. Al
   2. Pt
   3. Cu
   4. Zn

10. All of the following are used to prevent corrosion except
   1. painting the metal
   2. use of sacrificial metal
   3. coating with gold
   4. covering with oil

**True/False:**

11. _____ Hydrogen peroxide kills bacteria by oxidizing them.
12. _____ Oxidation is the removal of oxygen from a substance.
13. _____ Some reduction reactions do not involve oxygen.
14. _____ O in the water molecule is more electron-rich than H is.
15. _____ Zn is the oxidizing agent in the formation of ZnS from Zn and S.
16. _____ Net ionic equations omit oxidizing and reducing agents.
17. _____ The half-reaction \( \text{Cu}^{2+} \rightarrow \text{Cu} \) represents a reduction process.
18. _____ Hydrogen is oxidized when it combines with O to form water.
19. _____ Iron turns to rust when exposed to oxygen and water.
20. _____ Copper oxide reacts with water to form patina.

**Fill in the Blank:**

21. _______ involves a full or partial loss of ________ .
22. _______ involves a full or partial _____ of electrons.
23. A _______ is an equation that shows either the _______ or the reaction that occurs during a redox reaction.
24. Explain the role of a sacrificial metal in protection against corrosion.
25. For each of the following reactions, identify the entity that was oxidized and the entity that was reduced.

   Indicate the oxidizing and reducing agents:

   1. \( P_4 + 5O_2 \rightarrow P_4O_{10} \)
   2. \( 2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2 \)
   3. \( 2HgO \rightarrow Hg + O_2 \)
   4. \( 2NaH \rightarrow Na + H_2 \)
   5. \( Cl_2 + 2NaI \rightarrow 2NaCl + I_2 \)

26. Predict the products of the following reactions and write a balanced equation for each reaction:

   1. \( Br_2 + KI \)
   2. \( Mg + CuSO_4 \)
   3. octane \( (C_8H_{18}) + O_2 \)
4. $Al + Br_2$
5. $Cu + AgNO_3$

27. Write the reactions involved in the corrosion of iron.

### Answer Key

1. c  
2. d  
3. b  
4. b  
5. d  
6. b  
7. a  
8. d  
9. b  
10. c  
11. true  
12. false  
13. true  
14. true  
15. false  
16. false  
17. true  
18. true  
19. true  
20. false  
21. oxidation, electrons  
22. reduction, gain  
23. half-reaction, oxidation, reduction  
24. The sacrificial metal is one that is more active than the metal used in the construction. The sacrificial metal will undergo corrosion more rapidly and minimize the corrosion experienced by the other metal. Sacrificial metals are usually small blocks of metal attached to the construction and can easily be replaced.

#### Table 22.1:

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Oxidized</th>
<th>Reduced</th>
<th>Oxidizing Agent</th>
<th>Reducing Agent</th>
</tr>
</thead>
<tbody>
<tr>
<td>a</td>
<td>$P^0$ to $P^{+5}$</td>
<td>$O^{2-}$ to $O^{0}$</td>
<td>O</td>
<td>P</td>
</tr>
<tr>
<td>b</td>
<td>$Al^{0}$ to $Al^{3+}$</td>
<td>$H^+$ to $H^0$</td>
<td>H</td>
<td>$Al^{-}$</td>
</tr>
<tr>
<td>c</td>
<td>$O^{2-}$ to $O^{0}$</td>
<td>$Hg^{2+}$ to $Hg^{0}$</td>
<td>Hg</td>
<td>O</td>
</tr>
<tr>
<td>d</td>
<td>$H^{-1}$ to $H^{0}$</td>
<td>$Na^{+}$ to $Na^{0}$</td>
<td>Na</td>
<td>H</td>
</tr>
<tr>
<td>e</td>
<td>$I^{-1}$ to $I^{0}$</td>
<td>$Cl^{0}$ to $Cl^{-1}$</td>
<td>Cl</td>
<td>I</td>
</tr>
</tbody>
</table>

26. 1. $Br_2 + 2KI \rightarrow 2KBr + I_2$
2. $Mg + CuSO_4 \rightarrow Cu + MgSO_4$
3. $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$
4. $2Al + 3Br_2 \rightarrow 2AlBr_3$
5. $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$
27. $2Fe(s) + O_2(g) + 4H^+(aq) \rightarrow 2Fe^{2+}(aq) + 2H_2O(l)$
28. $4Fe^{2+}(aq) + O_2(g) + 6H_2O(l) \rightarrow 2Fe_2O_3 \cdot H_2O(s) + 8H^+(aq)$
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. The oxidation process involves all of the following except
   1. gain of oxygen
   2. loss of electrons
   3. gain of hydrogen
   4. increase in oxidation number

2. The reduction process involves all of the following except
   1. gain of hydrogen
   2. gain of oxygen
   3. gain of electrons
   4. decrease in oxidation number

3. The oxidation number for H in CaH₂ is
   1. +1
   2. +2
   3. -1
   4. -2

4. Oxygen has an oxidation number of -1 in
   1. Na₂O₂
   2. OF₂
   3. O₂
   4. MgO

5. The oxidation number for Br in LiBrO₃ is
   1. -1
   2. +2
   3. +4
   4. +5

6. In the reaction SF₄ + F₂ → SF₆, the oxidation change was
   1. S²⁺ → S⁴⁺
   2. F⁰ → F⁻⁶
   3. S⁴⁺ → S⁶⁺
   4. F⁺⁴ → F⁺⁶

7. One of the following is a type of redox reaction
   1. acid-base
   2. single-displacement
3. double-displacement  
4. neutralization  

8. In the reaction $KMnO_4 + HCl \rightarrow MnCl_2 + Cl_2 + H_2O + KCl$, all of the following statements are true except  
   1. K is oxidized  
   2. Cl is oxidized  
   3. Mn is reduced  
   4. H is unchanged  

9. One element for which there are no specific oxidation-reduction rules is  
   1. F  
   2. N  
   3. Ca  
   4. Fe  

10. Cl can have all of the following oxidation numbers except  
   1. -1  
   2. -2  
   3. +3  
   4. +5  

True/False:  

11. _____ In $K_3PO_4$, the sum of the oxidation numbers of the atoms equals zero.  
12. _____ The oxidation number for elemental silver is zero.  
13. _____ The oxidation number for S in $MgSO_4$ is +3.  
14. _____ The change in Cl for $HCl$ to $HClO_2$ represents an oxidation.  
15. _____ When NO is converted to NO$_2$, the nitrogen is reduced.  
16. _____ The highest possible oxidation number for N is +3.  
17. _____ The oxidation number for Ca$^{2+}$ is +2.  
18. _____ A free element is any element in an uncombined state.  
19. _____ A shift of electrons away from an atom in a covalent bond is a reduction of that atom.  
20. _____ Single-displacement reactions are redox reactions.  

Fill in the Blank:  

21. Overall, the _______number of an atom in a molecule is the _______that the atom would have if all polar ____ and ionic bonds resulted in a complete transfer of $s$ from the less electronegative atom to the more electronegative one.  
22. Which of the oxidation number rules is illustrated by each of the following:  
   1. $F^-$  
   2. $Pl^0$  
   3. $H^+$  
   4. $SO_4^{2-}$  

23. Which of the following are redox reactions?  
   1. $NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$  
   2. $HBr + NaOH \rightarrow NaBr + H_2O$  
   3. $C_{10}H_8 + 12O_2 \rightarrow 10CO_2 + 4H_2O$  
   4. $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$  
   5. $2HNO_3 + Sr(OH)_2 \rightarrow Sr(NO_3)_2 + 2H_2O$  
   6. $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$  

24. Indicate the oxidation number of the atoms listed
22.2. Oxidation Numbers

<table>
<thead>
<tr>
<th>Atom</th>
<th>Compound</th>
<th>Oxidation Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ba</td>
<td>Ba(NO(_3))(_2)</td>
<td></td>
</tr>
<tr>
<td>N</td>
<td>NF(_3)</td>
<td></td>
</tr>
<tr>
<td>S</td>
<td>(NH(_4))(_2)SO(_4)</td>
<td></td>
</tr>
<tr>
<td>Cr</td>
<td>K(_2)Cr(_2)O(_7)</td>
<td></td>
</tr>
</tbody>
</table>

25. For each of the following reactions, list the atom(s) oxidized and the atom(s) reduced:

1. \(8\text{NH}_3 + 6\text{NO}_2 \rightarrow 7\text{N}_2 + 12\text{H}_2\text{O}\)
2. \(\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2\)
3. \(5\text{CO} + \text{I}_2\text{O}_5 \rightarrow 5\text{CO}_2 + \text{I}_2\)
4. \(3\text{CuS} + 8\text{HNO}_3 \rightarrow 3\text{CuSO}_4 + 8\text{NO} + 4\text{H}_2\text{O}\)

**Answer Key**

1. c  
2. b  
3. c  
4. a  
5. d  
6. c  
7. b  
8. a  
9. b  
10. b 
11. true  
12. true  
13. false  
14. true  
15. false  
16. true  
17. true  
18. true  
19. false  
20. true  
21. oxidation, charge, covalent, electron
   1. rule 5  
   2. rule 1  
   3. rule 4  
   4. rule 7
22. reactions c, d, and f are redox reactions

<table>
<thead>
<tr>
<th>Atom</th>
<th>Compound</th>
<th>Oxidation Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ba</td>
<td>Ba(NO(_3))(_2)</td>
<td>+2</td>
</tr>
</tbody>
</table>
Table 22.3: (continued)

<table>
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<tr>
<th>Atom</th>
<th>Compound</th>
<th>Oxidation Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>N</td>
<td>NF₃</td>
<td>+3</td>
</tr>
<tr>
<td>S</td>
<td>(NH₄)₂SO₄</td>
<td>+6</td>
</tr>
<tr>
<td>Cr</td>
<td>K₂Cr₂O₇</td>
<td>+6</td>
</tr>
</tbody>
</table>
Lesson Quiz

The following questions deal with the oxidation number change method

1. The equation below illustrates one of the following steps

\[ ^{+2}Hg^{-2} \rightarrow ^0Hg + ^0O \]

2. \[ 2HgO \rightarrow 2Hg + O \]

   1. step 2
   2. step 1
   3. step 3
   4. step 4

3. ______ states that lines are used to connect atoms undergoing a change in oxidation number.

   1. step 1
   2. step 2
   3. step 3
   4. step 4

4. Checking to see if the equation is balanced is ______

   1. step 5
   2. step 6
   3. step 3
   4. step 2

5. Using coefficients to balance oxidation number changes is ______

   1. step 5
   2. step 7
   3. step 4
   4. step 3

6. _____ involves identifying oxidized and reduced atoms

   1. step 3
   2. step 2
   3. step 1
   4. step 4

The following questions deal with the half-reaction method.

6. \[ Al(s) + H^+ \rightarrow Al^{3+} + H_2(g) \] illustrates _____.

   1. step 2
   2. step 1
7. Balancing oxygen atoms by adding water where needed is ______
   1. step 4
   2. step 7
   3. step 3
   4. step 2

8. _____ calls for writing separate half reactions for oxidation and reduction.
   1. step 3
   2. step 2
   3. step 5
   4. step 4

9. Adding the two half-reactions together is _____
   1. step 3
   2. step 7
   3. step 6
   4. step 5

10. _____ calls for balancing charges by adding electrons
    1. step 4
    2. step 7
    3. step 3
    4. step 5

True/False:

11. _____ Balancing by inspection is not always successful.
12. _____ The oxidation number change method is more useful for reactions involving aqueous ions.
13. _____ Reactions in an acidic medium require the addition of protons.
14. _____ Coefficients are required for all components of a reaction.
15. _____ Oxidation of Fe$^{2+}$ by Cr$_2$O$_7^{2-}$ readily occurs at neural pH.
16. _____ The half-reaction process treats oxidation and reduction as two simultaneous processes.
17. _____ For reactions in basic solution, H$^+$ and OH$^-$ are combined to form water.
18. _____ Electrons do not need to balance on both sides of the equation.
19. _____ For reactions in basic solution, hydroxide ions are added.
20. _____ Most redox reactions can take place equally well in acidic and basic solutions.

21. Write half-reactions for the following processes:
    1. $Sn + NO_3^- \rightarrow SnO_2 + NO$
    2. $HClO + Co \rightarrow Cl_2 + Co^{2+}$
    3. $NO_2 \rightarrow NO_3^- + NO$

22. Balance the following equations:

23. Acidic Solution

   a. $Ag + NO_3^- \rightarrow Ag^+ + NO$
   b. $Zn + NO_3^- \rightarrow Zn^{2+} + NH_4^+$
   c. $Cr_2O_7^{2-} + C_2H_4O \rightarrow C_2H_4O_2 + Cr^{3+}$
   d. $H_3PO_2 + Cr_2O_7^{2-} \rightarrow H_3PO_4 + Cr^{3+}$
   e. Basic Solution
   f. $MnO_4^- + C_2O_4^{2-} \rightarrow MnO_2 + CO_2$
g. \( ClO^- + Fe(OH)_3 \rightarrow Cl^- + FeO_4^{2-} \)
h. \( HO_2^- + Cr(OH)_3 \rightarrow CrO_4^{2-} + OH^- \)
i. \( N_2H_4 + Cu(OH)_2 \rightarrow N_2 + Cu \)

**Answer Key**

1. b
2. c
3. a
4. c
5. b
6. b
7. a
8. b
9. c
10. d
11. true
12. false
13. true
14. true
15. false
16. true
17. true
18. false
19. true
20. false

1. \( Sn \rightarrow SnO_2 \) and \( NO_3^- \rightarrow NO_2 \)
2. \( HClO \rightarrow Cl_2 \) and \( Co \rightarrow Co^{2+} \)
3. \( NO_2 \rightarrow NO_3^- \) and \( NO_2 \rightarrow NO \)

1. \( 4H^+ + 3Ag + NO_3^- \rightarrow 3Ag^+ + NO + 2H_2O \)
2. \( 10H^+ + 4Zn + NO_3^- \rightarrow 4Zn^{2+} + NH_4^+ + 3H_2O \)
3. \( 8H^+ + Cr_2O_7^{2-} + 3C_2H_4O_2 \rightarrow 3C_2H_4O_2 + 2Cr^{3+} + 4H_2 \)
4. \( 16H^+ + 3H_3PO_2 + 2Cr_2O_7^{2-} \rightarrow 3H_3PO_4 + 4Cr^{3+} + 8H_2O \)
5. \( 4H_2O + 2MnO_4^- + 3C_2O_4^{2-} \rightarrow 2MnO_2 + 6CO_2 + 8OH^- \)
6. \( 4OH^- + 3ClO^- + 2Fe(OH)_3 \rightarrow 3Cl^- + 2FeO_4^{2-} + 5H_2O \)
7. \( OH^- + 3HO_2^- + 2Cr(OH)_3 \rightarrow 2CrO_4^{2-} + 5H_2O \)
8. \( N_2H_4 + 2Cu(OH)_2 \rightarrow N_2 + 2Cu + 4H_2O \)
Chapter Test

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. An example of an oxidation process is
   1. removal of oxygen
   2. reacting Fe₂O₃ with carbon
   3. heating mercuric oxide
   4. bleaching stains

2. An example of a reduction process is
   1. formation of MgO
   2. killing bacteria with H₂O₂
   3. refining iron ore
   4. rusting of iron

3. One metal that does not corrode easily is
   1. Fe
   2. Sn
   3. Au
   4. Mn

4. When O is reduced
   1. electrons are added to the O atom
   2. electrons are removed from the O atom
   3. electrons are transferred to another atom
   4. electrons are lost in the solution

5. The reduction process involves
   1. gain of oxygen
   2. loss of electrons
   3. gain of hydrogen
   4. increase in oxidation number

6. The oxidation process involves
   1. gain of hydrogen
   2. gain of oxygen
   3. gain of electrons
   4. decrease in oxidation number

7. One element for which there are no specific oxidation number rules is
   1. Na
   2. Ne
3. S
4. Ca

8. Br can have all of the following oxidation numbers except
   1. -1
   2. -2
   3. +3
   4. +5

9. The oxidation number for S in H₂SO₃ is
   1. -2
   2. -4
   3. +2
   4. +4

10. In the reaction 2KClO₃ → 2KCl + 3O₂, the oxidation change was
    1. Cl⁻ → Cl⁺⁵
    2. O²⁻ → O⁰
    3. Cl⁺³ → Cl⁻
    4. Cl⁺⁵ → Cl⁺³

True/False:

11. _____ When N₂O₃ changes to N₂O₅, the N has been oxidized
12. _____ The oxidation number for N in N₂F₄ is +3.
13. _____ Double-displacement reactions are redox reactions.
14. _____ O can have an oxidation number of -1 in some circumstances.
15. _____ O in the water molecule is less electron-rich than the H is.
16. _____ Net ionic equations eliminate spectator ions.
17. _____ Copper oxide reacts with CO₂ to form patina.
18. _____ The half-reaction Fe²⁺ → Fe³⁺ represents a reduction process.
19. _____ All oxidation reactions involve oxygen.
20. _____ A sacrificial metal helps prevent another metal from corrosion.

Fill in the Blank:

21. Define the following terms:
   1. oxidation-reduction reaction
   2. half-reaction
   3. corrosion
   4. oxidation number

22. Determine the oxidation number of the indicated element in each of the following compounds:

<table>
<thead>
<tr>
<th>Atom</th>
<th>Compound</th>
<th>Oxidation Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>H₂CO₃</td>
<td></td>
</tr>
<tr>
<td>Zn</td>
<td>Zn(OH)₄²⁻</td>
<td></td>
</tr>
<tr>
<td>H</td>
<td>LiH</td>
<td></td>
</tr>
<tr>
<td>Fe</td>
<td>Fe₂O₃</td>
<td></td>
</tr>
</tbody>
</table>
23. For each of the following reactions, identify the entity that was oxidized and the entity that was reduced. Indicate the oxidizing and reducing agents:

1. \(2Fe(OH)_2 + H_2O_2 \rightarrow 2Fe(OH)_3\)
2. \(2Cr^{3+} + H_2O + 6ClO_3^- \rightarrow Cr_2O_7^{2-} + 6ClO_2 + 2H^+\)
3. \(I_2O_5 + 5CO \rightarrow I_2 + 5CO_2\)

24. Predict the products of the following reaction and write balanced equations:

1. \(CuO + H_2 \rightarrow\)
2. \(KBr + Cl_2 \rightarrow\)
3. \(Fe + CuSO_4 \rightarrow\)

25. Balance the following equations:

1. \(MnO_4^- + I^- \rightarrow IO_3^- + Mn^{2+}\) under acid conditions.
2. balance the same equation under basic conditions
3. \(S_2O_3^{2-} + H_2O_2 \rightarrow S_4O_6^{2-} + H_2O\) under acid conditions

---

**Answer Key**

1. d
2. c
3. c
4. a
5. c
6. b
7. c
8. b
9. d
10. b
11. true
12. false
13. false
14. true
15. false
16. true
17. true
18. false
19. false
20. true

1. a reaction that involves the full or partial transfer of electrons from one reactant to another.
2. a reaction that involves the full or partial transfer of electrons from one reactant to another.
3. the deterioration of metals by redox processes.
4. a positive or negative number that is assigned to an atom to indicate its degree of oxidation or reduction.

---

**Table 22.5:**

<table>
<thead>
<tr>
<th>Atom</th>
<th>Compound</th>
<th>Oxidation Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>H_2CO_3</td>
<td>+4</td>
</tr>
<tr>
<td>Zn</td>
<td>Zn(OH)_4^{2-}</td>
<td>+2</td>
</tr>
</tbody>
</table>
### Table 22.5: (continued)

<table>
<thead>
<tr>
<th>Atom</th>
<th>Compound</th>
<th>Oxidation Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>LiH</td>
<td>-1</td>
</tr>
<tr>
<td>Fe</td>
<td>Fe₂O₃</td>
<td>+3</td>
</tr>
</tbody>
</table>

### Table 22.6:

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Oxidized</th>
<th>Reduced</th>
<th>Oxidizing Agent</th>
<th>Reducing Agent</th>
</tr>
</thead>
<tbody>
<tr>
<td>a</td>
<td>Fe +2 to +3</td>
<td>O -1 to -2</td>
<td>O</td>
<td>Fe</td>
</tr>
<tr>
<td>b</td>
<td>Cr 3⁺ to +6</td>
<td>Cl +5 to +4</td>
<td>Cl</td>
<td>Cr</td>
</tr>
<tr>
<td>c</td>
<td>C +2 to +4</td>
<td>I +5 to 0</td>
<td>I</td>
<td>C</td>
</tr>
</tbody>
</table>

1. \(\text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O}\)
2. \(2\text{KBr} + \text{Cl}_2 \rightarrow 2\text{KCl} + \text{Br}_2\)
3. \(\text{Fe} + \text{CuSO}_4 \rightarrow \text{FeSO}_4 + \text{Cu}\)

1. \(18\text{H}^+ + 6\text{MnO}_4^- + 5\text{I}^- \rightarrow 5\text{IO}_3^- + 6\text{Mn}^{2+} + 9\text{H}_2\text{O}\)
2. \(9\text{H}_2\text{O} + 6\text{MnO}_4^- + 5\text{I}^- \rightarrow 5\text{IO}_3^- + 6\text{Mn}^{2+} + 18\text{OH}^-\)
3. \(2\text{S}_2\text{O}_3^- + \text{H}_2\text{O}_2 + 2\text{H}^+ \rightarrow \text{S}_4\text{O}_6^{2-} + 2\text{H}_2\text{O}\)
Chapter Outline

23.1 Electrochemical Cells
23.2 Cell Potentials
23.3 Electrolysis
Lesson Quiz

Multiple Choice:

1. Electrochemistry involves all of the following applications except
   1. making batteries
   2. solar cells
   3. manufacturing metals
   4. refining aluminum

2. Writing separate half-reactions for a redox process is helpful for
   1. deciding which reaction occurs first
   2. predicting which half-reaction will not occur
   3. understanding the flow of electrons from one entity to another
   4. predicting alternate products

3. Tin can be oxidized by ions of
   1. copper
   2. magnesium
   3. nickel
   4. calcium

4. Calcium will not oxidize when in a solution of
   1. LiCl
   2. NiCl₂
   3. MgCl₂
   4. ZnCl₂

5. The first direct current cell was constructed by
   1. Le Châtelier
   2. Volta
   3. Röntgen
   4. Dalton

6. The notation for a voltaic cell comprised of copper and silver is
   1. Cu²⁺ | Cu | Ag⁺ | Ag
   2. Ag⁺ | Ag | Cu | Cu²⁺
   3. Cu | Cu²⁺ | Ag⁺ | Ag
   4. Cu | Cu²⁺ | Ag⁺ | Ag

7. In the Zn | Cu voltaic cell, the salt bridge is composed of
   1. ZnCl₂
   2. NaNO₃
3. CuCl₂
4. NaF

8. In a dry cell, the anode is
1. carbon
2. manganese dioxide
3. zinc
4. zinc chloride

9. An alkaline battery produces all of the following as products except
1. Mn(OH)₃
2. Zn(OH)₂
3. NH₃
4. H₂O

10. The anode in a lead storage battery is
1. PbO₂
2. PbO
3. Pb
4. PbSO₄

**True/False:**

11. _____ Oxidation and reduction reactions occur at the same time.
12. _____ Al will spontaneously oxidize in a solution of Ni(NO₃)₂.
13. _____ Lead will spontaneously oxidize in a solution of NaCl.
14. _____ Some spontaneous reactions can produce an electrical current.
15. _____ The two electrodes are connected by a salt bridge.
16. _____ Oxidation takes place at the anode.
17. _____ The electrolyte in the salt bridge is often AgNO₃.
18. _____ The cathode for a dry cell battery is carbon.
19. _____ In a dry cell battery, the paste eliminates the need for a salt bridge.
20. _____ An alkaline battery contains a paste of Zn and KOH.

**Fill In the Blank:**

21. _______ is a branch of chemistry that deals with the _________of chemical energy and electrical energy.
22. An element that is _______in the activity series is capable of _______an element that is lower on the series in a ______reaction.
23. An _______ cell is any device that converts _______energy into electrical energy or electrical energy into chemical energy.
24. An _______is a conductor in a circuit that is used to carry to a ______part of the circuit.
25. In each of the following combinations, indicate whether or not a reaction will occur:
   1. Fe(s) + CaSO₄(aq)
   2. Ba(s) + AgNO₃(aq)
   3. SnCl₂(aq) + Mg(s)
   4. CaCl₂(aq) + Ni(s)

**Short Answers:**

26. Write the shorthand notation for the cell that uses the reaction
27. Al(s) + Sn²⁺(aq) → Al³⁺(aq) + Sn(s)
23.1. Electrochemical Cells

28. Can a lead storage battery recharge completely? Explain your answer.
29. What is one safety problem with a hydrogen-oxygen fuel cell?
30. A cell has the following diagram:
31. \( Ba^{2+} | | Ba | Cu | Cu^{2+} \)
32. Write the two half-reactions for this cell

---

**Answer Key**

1. b
2. c
3. a
4. a
5. b
6. c
7. b
8. c
9. a
10. c
11. true
12. true
13. false
14. true
15. true
16. true
17. false
18. false
19. false
20. true
21. electrochemistry, interconversion
22. higher, displacing, single-replacement
23. electrochemical, chemical
24. electrode, electrons, nonmetallic
   1. no reaction
   2. reaction
   3. no reaction
   4. no reaction
25. \( Al | Al^{3+} || Sn^{2+} | Sn \)
26. No, the lead sulfate that is formed cannot regenerate both metallic lead and lead oxide to any great extent.
27. The fuel for this cell is hydrogen gas, which is a very flammable material. Delivery of this gas to a heated system for reaction with oxygen is very complicated and storage is a challenge.
28. \( Ba \rightarrow 2e^- + Ba^{2+} \)
29. \( Cu^{2+} + 2e^- \rightarrow Cu \)
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. Movement of electrons is responsible for the _______ of a cell
   1. electrical charge
   2. electrical potential
   3. electrical difference
   4. electrical partition

2. Half-cell processes produce
   1. resistance
   2. milliamps
   3. voltage
   4. ohms

3. Standard cell potentials are measured in
   1. kilovolts
   2. millivolts
   3. microvolts
   4. volts

4. The standard condition for cells using gases is
   1. 700 mm Hg
   2. 800 mm Hg
   3. 760 mm Hg
   4. 740 mm Hg

5. A positive standard reduction potential means
   1. it is easier to reduce the species
   2. it is easier to reduce hydrogen
   3. it is easier to oxidize the species
   4. hydrogen oxidizes less readily than the species

6. The half-cell with the lower reduction potential will
   1. undergo reduction
   2. undergo oxidation
   3. cause the other cell to oxidize
   4. gain electrons

7. Cd$^{2+}$ will undergo reduction when paired with a half-cell made of
   1. Fe$^{3+}$
   2. Cu$^+$
3. Fe^{2+}
4. Sn^{4+}

8. A strong reducing agent is one that can
   1. pull electrons away from other species
   2. donate electrons to other species
   3. transfer electrons from other species
   4. convert metals to metallic ions

9. Cu^{2+} will oxidize all of the following except
   1. Pb
   2. Sn
   3. Cr
   4. Ag

10. $E_{\text{red}} - E_{\text{ox}}$ is the formula for calculating
    1. cell potential
    2. cell current
    3. cell resistance
    4. cell activation

True/False:
11. The potential of an isolated half-cell cannot be calculated.
12. Electrical energy is produced by double-displacement reactions.
13. The standard hydrogen electrode is immersed in 0.1 M H^+.
14. The standard hydrogen electrode is made of platinum.
15. H_2 is formed when the standard hydrogen electrode is involved in a reduction reaction.
16. A redox reaction occurs when the cell potential is negative.
17. I_2 is more readily reduced than Ni^{2+}
18. Ag^{3+} reduces more readily than Hg^{2+}
19. In a cell, the half-cell with the higher potential will undergo oxidation.
20. Lithium is the strongest reducing agent.

Fill in the Blank:
21. Electrical ______ is a measurement of the ability of a _______ cell to produce an electric current.
22. The cell potential ($E_{\text{cell}}$) is the difference in ________ potential between the two half-cells in an _________-_ cell.

Short Answers:
23. State the conditions for measuring standard cell potentials.
24. In the hydrogen electrode, when is hydrogen gas produced and when are hydrogen ions formed?
25. Which of the following metals will react with HCl? Ag, Zn, Cu, Ni, Fe?
26. Which of the following is the strongest oxidizing agent? The weakest?
27. Cr, Cu, Zn, Al, Au
28. In a typical half-cell, a metal strip is immersed in a solution containing the cation of that metal. What fulfills these functions in the hydrogen half-cell?
29. Which of the following reactions will occur spontaneously?
   1. $Co + Sn^{4+} \rightarrow Co^{2+} + Sn$
   2. $Mn + Zn^{2+} \rightarrow Mn^{2+} + Zn$
3. $Ag + Cu^{2+} \rightarrow Ag^+ + Cu$

30. Calculate the standard cell potential produced by a galvanic cell consisting of a nickel electrode in contact with a solution of $Ni^{2+}$ ions and a silver electrode in contact with a solution of $Ag^+$ ions. Which is the anode and which is the cathode?

31. What is the voltage produced by a galvanic cell consisting of an aluminum electrode in contact with a solution of $Al^{3+}$ ions and an iron electrode in contact with a solution of $Fe^{2+}$ ions. Which is the anode and which is the cathode?

---

**Answer Key**

1. b
2. c
3. d
4. c
5. a
6. b
7. c
8. b
9. d
10. a
11. true
12. false
13. false
14. true
15. true
16. false
17. true
18. true
19. false
20. true
21. potential, voltaic
22. reduction, electrochemical
23. temperature: 25°C, all aqueous components at a concentration of 1 M, all gases at 1 atm pressure.
24. Hydrogen gas is produced when the hydrogen half-cell is connected to a half cell with a lower reduction potential. Hydrogen ions are formed when the hydrogen electrode is connected to a half-cell with a higher reduction potential.
25. Zn, Ni
26. strongest: Al weakest Au
27. In the hydrogen half-cell, $H_2$ gas is bubbled into the system and takes the place of the metal. Hydrogen ions in solution serve the role of the cation for the half-cell.

1. Spontaneous
2. Spontaneous
3. Nonspontaneous

28. Ni is below Ag in the table, so it will be oxidized and $Ag^+$ will be reduced
29. $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s) \quad E^o = -0.25 \, V \quad Ag^+(aq) + e^- \rightarrow Ag(s) \quad E^o = 0.80 \, V$
30.

Anode: \[ \text{Ni}(s) \rightarrow \text{Ni}^{2+}(aq) + 2e^- \quad E^\circ = 0.25 \text{ V} \text{ (Sign reversed for oxidation)} \]

Cathode: \[ \text{Ag}^+(aq) + e^- \rightarrow \text{Ag}(s) \quad E^\circ = 0.80 \text{ V} \]

\[ E^\circ = 1.05 \text{ V} \]

31. Al is below iron in the table so it will be oxidized and iron will be reduced

32. \[ \text{Al}^{3+}(aq) + 3e^- \rightarrow \text{Al}(s) \quad E^\circ = -1.66 \text{ V} \quad \text{Fe}^{2+}(aq) + 2e^- \rightarrow \text{Fe}(s) \quad E^\circ = -0.44 \text{ V} \]

33.

Anode: \[ \text{Al}(s) \rightarrow \text{Al}^{3+}(aq) + 3e^- \quad E^\circ = 1.66 \text{ V} \text{ (Sign reversed for oxidation)} \]

Cathode: \[ \text{Fe}^{2+}(aq) + 2e^- \rightarrow \text{Fe}(s) \quad E^\circ = -0.44 \text{ V} \]

\[ E^\circ = 1.22 \text{ V} \]
Lesson Quiz

1. Electrolysis is used for
   1. spontaneous reactions
   2. nonspontaneous reactions
   3. reversible reactions
   4. double-displacement reactions

2. In an electrolytic reaction involving Pb and Ni
   1. Pb is reduced
   2. Ni is reduced
   3. Pb is the anode
   4. Ni is the cathode

3. The Down’s cell uses _____ to produce its products
   1. molten Na
   2. gaseous Cl
   3. molten NaCl
   4. aqueous NaCl

4. The cell potential for the anode reaction in the Down’s cell is
   1. -1.36 V
   2. -2.71 V
   3. +1.36 V
   4. +2.71 V

5. The electrode used in the apparatus for the electrolysis of water is
   1. Au
   2. Fe
   3. Ag
   4. Pt

6. The electrolyte used for the electrolysis of water is
   1. HCl
   2. H₂SO₄
   3. H₃PO₄
   4. HClO

7. Na metal is not produced in the electrolysis of brine because
   1. the reduction potential for water is more negative
   2. the reduction potential for Na is less negative
3. the reduction potential for water is less negative
4. the reduction potential for water is more positive

8. One of the following statements about the brine electrolysis is true
   1. the reduction reaction forms hydrogen gas and hydroxide ion
   2. the reduction process forms oxygen gas and hydroxide ion
   3. the oxidation process forms chloride ions
   4. the oxidation process produces hydroxide ions

9. Electroplating with gold would involve all of the following except
   1. oxidizing gold metal from the anode
   2. oxidizing gold metal from the cathode
   3. reducing gold ions at the cathode
   4. using a solution of gold nitrate

10. One metal that would not be safe for electroplating is
    1. Fe
    2. Cu
    3. Hg
    4. Ni

True/False:

11. Platinum is commonly plated onto objects
12. A voltaic cell uses an electric current to initiate a nonspontaneous reaction
13. NaOH is a major product of the Down’s cell.
14. Aqueous NaCl is used in the Down’s cell.
15. Both O₂ and H₂ are formed during the electrolysis of water.
16. In electrolysis, OH⁻ is formed at the anode.
17. During the electrolysis of brine, hydrogen gas is produced at the cathode.
18. NaOH is an important product of brine electrolysis.
19. In electroplating, silver is oxidized at the cathode.
20. Silver ions for electroplating come from the silver nitrate solution.

Fill in the Blank:

21. If a chemical system is supplied with ______ from an external source, it is possible to drive a reaction in the direction.
22. is the process in which ______ energy is used to cause a chemical reaction to occur.

Short Answers:

23. If a voltaic cell has an arrangement in which the Zn electrode is the anode and the Cu electrode is the cathode, what will be the arrangement in an electrolytic cell?
24. What will the standard cell potential be for a nonspontaneous reaction?
25. What is the direction of electron flow in an electrolytic cell?
26. Why does sulfuric acid need to be added to water for electrolysis?
27. What would happen if HCl were used in the electrolysis of water?
28. Explain the process taking place in the following diagram. Me stands for a metal:
29.
30. Explain the process taking place in the following diagram:

31.

Answer Key

1. b
2. c
3. a
4. a
23.3. Electrolysis

5. d
6. b
7. c
8. a
9. b
10. c
11. true
12. true
13. false
14. false
15. true
16. false
17. true
18. true
19. false
20. true
21. energy, nonspontaneous
22. electrolysis, electrical, nonspontaneous
23. Zn will now be the cathode and Cu the anode.
24. negative
25. from anode to cathode
26. pure water does not conduct electricity well
27. It is possible that the chloride ions could be reduced to chlorine gas.
28. The metal (identity unknown) is being plated by the copper. At the anode, copper metal is being oxidized to \( \text{Cu}^{2+} \). At the cathode, the copper ions are being reduced to metallic copper which then plates out onto the other metal.
29. Fused (melted) \( \text{NaCl} \) enters the cell. The carbon anode oxidized the \( \text{Cl}^- \) to form \( \text{Cl}_2 \) gas which leaves through the center vent at the top of the diagram. The iron cathode reduces the \( \text{Na}^+ \), forming liquid sodium metal. The sodium metal is less dense than the molten \( \text{NaCl} \) and floats to the top of its trap (left hand side of diagram).
Chapter 24

Nuclear Chemistry Assessments

Chapter Outline

24.1 Nuclear Radiation
24.2 Half Life
24.3 Fission and Fusion
24.4 Nuclear Chemistry
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice:

1. Lithium-6 has a measured mass of 6.015121 amu. This value
   1. equals the mass of 3 protons and 3 neutrons
   2. equals the mass of three protons, three neutrons, and three electrons
   3. is more than the mass of three protons, three neutrons, and three electrons
   4. is less than the mass of three protons, three neutrons, and three electrons

2. Using Einstein’s famous equation the mass defect can be used to calculate
   1. nuclear energy
   2. nuclear binding energy
   3. nuclear stabilizing energy
   4. nuclear attractive energy

3. The force holding protons together is called the
   1. electrostatic force
   2. nuclear force
   3. strong nuclear force
   4. strong electrostatic force

4. No stable nuclei are seen beginning with element
   1. 100 –fermium
   2. 90 –thorium
   3. 83 –bismuth
   4. 93 –neptunium

5. Neutrons are needed to overcome the ______ of the protons in the nucleus
   1. electrostatic repulsion
   2. electrostatic force
   3. repulsive force
   4. electrostatic attraction

6. The discovery of radioactivity disproved one of the assumptions of _____ about atoms
   1. Amedeo Avogadro
   2. Robert Boyle
   3. Antoine Lavoisier
   4. Jon Dalton

7. A radioactive event produces an atom with an atomic mass two amu less than its parent. The particle emitted was
   1. a positron
2. an alpha particle
3. a beta particle
4. gamma emission

8. On radioactive decay, the daughter nuclide has an atomic number one less than the parent nuclide. The particle emitted was
   1. a positron
   2. an alpha particle
   3. a beta particle
   4. gamma emission

9. On radioactive decay, the daughter nuclide has an atomic number one greater than the parent nuclide. The particle emitted was
   1. a positron
   2. an alpha particle
   3. a beta particle
   4. gamma emission

10. Electron capture by a nucleus produces the following:
    1. atomic number increased by one
    2. atomic number decreased by one
    3. atomic mass increased by one
    4. atomic mass decreased by one

**True/False**

11. _____ Temperature and pressure affect the rate of nuclear reactions.
12. _____ Uranium and plutonium have no stable isotopes.
13. _____ Uranium salts fog photographic plates only after being exposed to sunlight.
14. _____ The nucleus of an atom increases in stability after a radioactive event.
15. _____ Electrons help provide stability to nuclei.
16. _____ Stable nuclei with odd numbers of protons and neutrons are more common than nuclei with other combinations of protons and neutrons.
17. _____ Gamma radiation decreases the atomic number by one.
18. _____ Alpha particles have low penetrating power.
19. _____ A beta particle is a high-speed neutron emitted during a decay process.
20. _____ Gamma rays have very high penetrating power.

**Fill in the Blank**

21. ________ is the spontaneous breakdown of an atom’s ________ by the emission of ________ and/or radiation.
22. A ________is an isotope of an element that is ________and undergoes radioactive______.
23. Where does the mass go that is related to the mass defect?
24. Calculate the binding energy of the nuclide \( ^{10}_5B \) where the mass of \( ^{10}_5B \) atom = 10.0129 amu and the mass defect is \( 1.1093\times10^{-28} \) Kg
25. \( ^{238}_{92}U \) decays into \( ^{234}_{90}Th \) and an alpha particle.
   1. Write down the full decay equation
   2. How much energy is released.

    Mass of \( ^{238}_{92}U = 238.0508u \)
    Mass of \( ^{234}_{90}Th = 234.0426u \)
    Mass of \( ^{3}_2\alpha = 4.0026u \)
26. Plans are made to build a facility for studying radioisotope decay reactions. Concerns are expressed about exposure to radiation. You are assured that the lab will be surrounded with five layers of aluminum foil to protect anyone in another room. Is this adequate protection? Explain your answer.

27. Complete the following reactions:

1. $^{231}_{91}Pa \rightarrow ^4_2He + ?$
2. $? \rightarrow ^\beta + ^{223}_{88}Ra$
3. $^{201}_{80}Hg + ? \rightarrow ^{201}_{79}Au$
4. $^{120}_{55}Cs \rightarrow \text{positron} + ?$

Answer Key

1. d
2. b
3. c
4. c
5. b
6. d
7. b
8. a
9. c
10. a
11. false
12. true
13. false
14. true
15. false
16. false
17. false
18. true
19. false
20. true
21. Radioactivity, nucleus, particles
22. radioisotope, unstable, decay
23. That mass is converted to energy and serves as binding energy for the nucleus.
24. $E = mc^2 = 1.1093 \times 10^{-28} \text{ kg} \times (3 \times 10^8 \text{ m/s})^2 = 9.9836 \times 10^{-12} \text{ J}$
   a. $^{238}_{92}U \rightarrow ^{234}_{90}Th + ^4_2\alpha$
   b. First calculate mass change
      
      238.0508 amu - (234.0426 amu + 4.0026 amu)
      mass change = $5.6 \times 10^{-3}$ amu
      Convert to kg: $5.6 \times 10^{-3}$ amu $\times \frac{g}{6.022 \times 10^{23} \text{ amu}} \times \frac{kg}{1000 g} = 9.2988 \times 10^{-30} kg$
      Energy released $E = mc^2 = 9.2988 \times 10^{-30} \text{ kg} \times (3 \times 10^8 \text{ m/s})^2 = 8.36892 \times 10^{-13} \text{ J}$
25. This is not adequate protection. Gamma emissions can penetrate many materials, but will be blocked by lead shielding. Aluminum foil will not provide adequate protection.

1. $^{227}_{89}Ac$
2. $^{223}_{87}Fr$
3. electron capture
4. $^{120}_{54}Xe$
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice

1. The rate of decay depends on
   1. size of sample
   2. half-life of isotope
   3. temperature
   4. pressure

2. After three half-lives, _____ of the isotope remains
   1. 12.5 %
   2. 6.25 %
   3. 25 %
   4. 3.125 %

3. The age of the earth has been assessed using ______ isotopes
   1. radium
   2. plutonium
   3. carbon
   4. uranium

4. In the decay scheme isotope A → isotope B, isotope A is referred to as the
   1. daughter nuclide
   2. parent nuclide
   3. mother nuclide
   4. reactant nuclide

5. U-238 first decays to
   1. Pa-234
   2. U-234
   3. Th-234
   4. U-236

6. Pb-206 is formed from
   1. Po-210
   2. Bi-210
   3. Pb-208
   4. Pb-210

7. Fermium is a ______ element
   1. alkaline earth
   2. alkali
3. transuranium
4. transnucleic

8. Neutron bombardment produces all of the following except
   1. alpha particles
   2. beta particles
   3. neutrons
   4. positrons

9. Carbon-14 is used for
   1. dating organisms
   2. dating rocks
   3. dating seawater
   4. dating atmospheric CO₂

10. Rutherford bombarded nitrogen gas with alpha particles to produce
    1. F-20
    2. Ne-21
    3. F-18
    4. Ne-19

**True/False**

11. _____ Fr-220 has a half-life of 12.4 years.
12. _____ The abbreviation for half-life is \( t_{1/2} \).
13. _____ Cobalt-60 decays by alpha emission.
14. _____ Dubnium has been formed by artificial transmutation.
15. _____ The rate of decay of P-32 is enhanced by heating.
16. _____ The ratio of C-14 to C-12 is constant while the organism is alive.
17. _____ Artificial transmutation occurs spontaneously.
18. _____ James Chadwick discovered the neutron in 1932.
19. _____ Technetium still occurs in nature.
20. _____ Radioactive dating gives precise information about the age of an object.

**Fill in the Blank**

21. Why does the ratio of C-14 to C-12 change after the organism has died?
22. ______ transmutation is the bombardment of _____ nuclei with charged or uncharged particles in order to cause a ______ reaction.
23. N-16 has a half-life of 7.2 seconds. How much of the original activity is left after 28.8 seconds?
24. Sr-90 has a half-life of 28.1 days. If the original sample had a mass of 12.4 grams, how much would remain three half-lives later?
25. A sample of a radioisotope had a mass of 40.0 g at 12:30 PM. The material had a mass of 5 g at 1:20 PM. What is the half-life?
26. Complete the following reactions:
   1. \(^{48}\text{Ti} + \alpha \rightarrow p + ?
   2. \(^{62}\text{Cu} + n \rightarrow ^{60}\text{Co} + ?
   3. \(^{40}\text{Ca} + ? \rightarrow ^{47}\text{Sc} + p
   4. \(^{9}\text{Be} + \alpha \rightarrow n + ?

27. Write the shorthand notation for the following reactions:
   1. \(^{14}\text{N} + ^{1}\text{H} \rightarrow ^{17}\text{O} + ^{1}\text{H}
   2. \(^{9}\text{Be} + ^{1}\text{H} \rightarrow ^{6}\text{Li} + ^{2}\text{He}

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Answer Key

1. b
2. a
3. d
4. b
5. c
6. a
7. c
8. d
9. a
10. c
11. false
12. true
13. false
14. true
15. false
16. true
17. false
18. true
19. false
20. false
21. There is no more intake from the environment, so there is no C-14 being added to the organism’s carbon pool.
22. Artificial, stable, nuclear
23. activity = \((1/2)^4\) = 6.25\% of original activity.
24. \((12.4 \text{ g})(1/2)^3 = (12.4 \text{ g})(0.125) = 1.55 \text{ g}\)
25. Activity loss: in one half-life, the material goes to 20 g, two half-lives 10 g, three half-lives 5 g. So there are three half-lives in 50 minutes.
   
   50 min \div 3 \text{ half-lives} = 16.7 \text{ minutes/half-life.}

1. \(^{50}\text{V}\)
2. \(\alpha\)
3. \(\alpha\)
4. \(^{12}\text{C}\)

1. \(^{14}\text{N}(\alpha, p)^{17}\text{O}\)
2. \(^9\text{Be}(p, \alpha)^3\text{Li}\)
24.3 Fission and Fusion

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice

1. A chain reaction is usually triggered by
   1. a proton
   2. a neutron
   3. an electron
   4. a positron

2. When U-235 is hit by a neutron, it temporarily becomes
   1. U-236
   2. U-238
   3. Pu-236
   4. Pu-240

3. The primary radioactive material used in a nuclear reactor is
   1. plutonium
   2. curium
   3. uranium
   4. beryllium

4. Electric power in a nuclear reactor is produced by
   1. bombardment of control rods
   2. generation of steam
   3. high-energy neutrons
   4. nuclear fusion

5. The explosion in an atomic bomb is triggered by compression of
   1. uranium
   2. polonium
   3. plutonium
   4. beryllium

6. Energy in the sun is generated by fusion of
   1. hydrogen nuclei
   2. helium nuclei
   3. lithium nuclei
   4. beryllium nuclei

7. All of the following are advantages of fusion energy except
   1. no radioactive waste produced
   2. extremely high energies available
3. only reactant is hydrogen
4. very high temperatures involved

8. Radiation produced by radioisotopes is called
   1. ionization
   2. ionizing radiation
   3. ionized radiation
   4. ionizable radiation

   1. a gas-filled tube
   2. a vacuum tube
   3. a water-filled tube
   4. a diode tube

10. The isotope used to detect thyroid problems is
    1. I-129
    2. P-32
    3. I-131
    4. P-34

True/False

11. _____ A fusion process will produce more energy than a fission process.
12. _____ Solar power involves the use of radioisotopes to generate electricity.
13. _____ U-235 splits into Xe and Be when struck by a neutron.
14. _____ Control rods absorb some of the neutrons in a nuclear reactor.
15. _____ Fusion takes place at extremely high temperatures.
16. _____ Alpha particles can be stopped by skin and clothing.
17. _____ Background radiation exposure is over 2.0 rem/year for an individual.
18. _____ A scintillation counter gives off flashes of light when ionizing radiation is present.
19. _____ Co-58 is used to treat cancer.
20. _____ Ionizing radiation can be used to kill bacteria that cause food spoilage.

Fill in the Blank

21. Nuclear ________ is a process in which a very _______ nucleus (mass >200) splits into _______ nuclei of intermediate mass.
22. Control rods limit the amount of ______ ________ by absorbing some of them and preventing the _______-
    from proceeding too ________.
23. ______ radiation is radiation that has enough energy to knock _______ off the atoms of a bombarded
    substance and produce ions.
24. What would be the charge of the ions produced by the process in question 23?
25. What are the advantages and disadvantages of nuclear power when compared to power generated by fossil
    fuels?
26. Predict the products of a fusion reaction between deuterium (H-2) and tritium (H-3): $^2_1H + ^3_1H \rightarrow$?
27. Why are control rods needed in nuclear reactors?
28. One of the reactions that takes place in the sun is the collision of two $^3_2He$ particles to form an atom of $^4_2He$.
    What other two particles are also released?
29. A solar fusion reaction involves the collision of N-15 and a proton. One of the products is He-4. What is the other
    product?
30. A collision between U-235 and a neutron produces $^{94}_{38}Sr +$ two neutrons. What is the other product produced
    in this reaction?


Answer Key

1. b
2. a
3. c
4. b
5. c
6. a
7. d
8. b
9. a
10. c
11. true
12. false
13. false
14. true
15. true
16. true
17. false
18. true
19. false
20. true
21. fission, heavy, smaller
22. available neutrons, reaction, rapidly
23. Ionizing, electrons
24. If electrons are dislodged, the remaining particle has less electrons and would have a positive charge.
25. Advantages: no production of carbon dioxide or other air pollutants.
   Disadvantages: nuclear waste is radioactive and must be stored or disposed of properly. An accident at a nuclear power plant can be life-threatening.
26. $^4_2He + n$
27. To slow down high-speed neutrons and to moderate the number of neutrons available for further reaction.
28. Two protons. $2$ He-3 (p+p+n) → He-4 (2p, 2n) + 2p
29. C-12.
30. $^{140}_{54}$Xe
24.4 Nuclear Chemistry

Chapter Test

Name _____________________ Class ______________________ Date ________________

Multiple Choice

1. Nuclear binding energy is calculated using
   1. $E = m \cdot c^2$
   2. $E = mc^2$
   3. $E = c^2/m$
   4. $E = m/c^2$

2. The nuclear decay of C-14 produces N-14 and
   1. a positron
   2. an alpha particle
   3. a beta particle
   4. gamma emission

3. The mass of Fe-56 is 55.934937 amu and the mass defect is 0.528479 amu. This means that
   1. the measured mass is less than the sum of protons + neutrons
   2. the measured mass is greater than the sum of protons + neutrons
   3. the measured mass is greater than the sum of protons + neutrons + electrons
   4. the measured mass is less than the sum of protons + neutrons + electrons

4. undergoes electron capture to yield
   1. Te-125
   2. Te-123
   3. Xe-131
   4. Xe-129

5. Radium is a _____ element
   1. alkaline earth
   2. alkali
   3. transuranium
   4. transnucleic

6. Uranium is used to date
   1. organic material
   2. water samples
   3. atmospheric contamination
   4. age of rocks and minerals

7. The predominant particle emitted when U-238 decays is
   1. a positron
   2. an alpha particle
3. a beta particle
4. gamma emission

8. Slow neutrons are involved in all of the following except
   1. nuclear reactors
   2. U-235 fission
   3. hydrogen fusion
   4. production of transuranium elements

9. A gas-filled tube is used in a ____ counter to detect radioactivity.
   1. Müller
   2. scintillation
   3. isotope
   4. Geiger

10. Plutonium is used in
    1. atomic bombs
    2. nuclear reactors
    3. food sterilization
    4. fertilizer studies

True/False

11. _____ Francium and radium have no stable isotopes.
12. _____ Stable nuclei with even numbers of protons and neutrons are more common than nuclei with other combinations of protons and neutrons.
13. _____ A positron emission increases the atomic number by one.
14. _____ The electrostatic repulsions in the nucleus are overcome by neutrons.
15. _____ Gamma radiation increases the atomic number by one.
16. _____ Cobalt-60 decays by positron emission.
17. _____ Bohrium has been produced by artificial transmutation.
18. _____ A fission process produces more energy than a fusion process.
19. _____ Beta particles can be stopped by skin and clothes.
20. _____ I-131 is used to diagnose and treat thyroid cancer.

21. Define the following terms:
1. nucleon
2. radioactive decay
3. nuclear fission
4. nuclear fusion

22. The mass defect for Cu-63 is 9.8346*10^{-28} kg. Calculate the nuclear binding energy.

23. Complete the following reactions:
1. \( ^{222}_{86} Rn \rightarrow ^{218}_{84} Po + ? \)
2. \( ? + \) electron capture \( \rightarrow ^{116}_{49} \) In
3. \( \frac{64}{29} Cu \rightarrow \) positron + \(? \)
4. \( \frac{40}{19} K \rightarrow ^{40}_{20} \) Ca + ?

24. P-32 has a half-life of 14.3 days. How much of the original activity will remain after 85.8 days?
25. Th-234 has a half-life of 24.1 days. A sample has an initial mass of 23.7 grams. How much will be left after 72.3 days?

26. Complete the following equations:
1. \( \frac{56}{26} Fe + \frac{1}{1} \rightarrow ^{54}_{25} Mn + ? \)
2. \( \frac{106}{46} Pd + \alpha \rightarrow p + ? \)
Answer Key

1. b
2. c
3. d
4. b
5. a
6. d
7. b
8. c
9. d
10. a
11. false
12. true
13. false
14. true
15. false
16. false
17. true
18. false
19. false
20. true

1. a nuclear subatomic particle (either a proton or a neutron).
2. a reaction in which a nucleus spontaneously disintegrates into a slightly lighter nucleus, accompanied by
   the emission of particles, energy, or both.
3. is a process in which a very heavy nucleus (mass >200) splits into smaller nuclei of intermediate mass.
4. is a process in which light-mass nuclei combine to form a heavier and more stable nucleus.

21. \[ E = \left(9.8346 \times 10^{-28} \text{ kg/nucleus}\right) \left(3.0 \times 10^{8} \text{ m/s}\right)^2 = 8.85 \times 10^{-11} \text{ J} \]

   1. alpha particle
   2. $^{116}_{50} Sn$
   3. $^{61}_{28} Ni$
   4. beta particle

22. 85.8 days ÷ 14.3 days/half-life = 6 half-lives.
    100% x (1/2)$^6$ = 100 x 1/64 = 1.56%

23. 72.3 days ÷ 24.1 days/half-life = 3 half-lives
    23.7 g x (1/2)$^3$ = 23.7g x 1/16 = 1.48 g

   1. alpha particle
   2. $^{109}_{47} Ag$
Chapter 25

Organic Chemistry Assessments

Chapter Outline

25.1 HYDROCARBONS
25.2 FUNCTIONAL GROUPS
25.3 ORGANIC REACTIONS
25.4 ORGANIC CHEMISTRY
Lesson Quiz

1. The bond angles in cyclopropane are
   1. 60°
   2. 108°
   3. 109.5°
   4. 90°

2. The most stable cycloalkane structure is
   1. cyclohexane
   2. cyclopropane
   3. cyclopentane
   4. cyclobutane

3. The most stable cyclohexane ring conformation is the
   1. chair
   2. boat
   3. square
   4. stand

4. A cycloalkene has
   1. at least one double bond in the ring
   2. at least one triple bond in the ring
   3. at least one double bond on a side-chain substituent
   4. at least one triple bond on a side-chain substituent

5. The benzene ring structure was proposed by
   1. the German chemist Alexander Kekulé
   2. the French chemist August Châtelier
   3. the German chemist August Kekulé
   4. the German chemist August Châtelier

6. The bonding in the benzene ring is an example of
   1. delocalized pi bonding
   2. pi bonding
   3. delocalized sigma bonding
   4. sigma bonding

7. The benzene molecule is often represented by
   1. a six-membered ring
   2. a circle
3. a ring inside a circle
4. a circle inside a ring

8. An ortho substituent on a benzene ring means that there are ____ attached to the ring.
   1. three groups
   2. four groups
   3. two groups
   4. one group

9. A benzene ring that serves as a substituent on a carbon chain is known as a ______ group.
   1. benzyl
   2. phenyl
   3. phenol
   4. benzoic

10. The benzene ring has a ______ configuration
    1. puckered
    2. boat
    3. planar
    4. chair

True/False

11. _____ Cyclic hydrocarbons can be represented by ring structures.
12. _____ Cycloalkanes have the general formula C\(_n\)H\(_{2n+2}\).
13. _____ Cyclooctane has eight sides.
14. _____ Carbon atoms in cycloalkanes are sp\(^3\) hybridized.
15. _____ Cyclohexene has two double bonds in the ring.
16. _____ Benzene is more reactive than cyclohexene.
17. _____ Pi bonding electrons in benzene are mobile.
18. _____ Vanillin is a cyclohexene derivative.
19. _____ Chlorobenzene has a Cl atom attached to the benzene ring.
20. _____ Naming of benzene substituents is done by number, not alphabetically.

Fill in the Blank

21. Identify each of the following as a cycloalkane, cycloalkene, or benzene derivative:
22. Name the following compounds:
23. Draw the following structures:
   a. 1,2-dibromocyclopentane
   b. 1-chloro-3-iodobenzene
   c. 1-butyl-4-ethylbenzene
   d. 2-ethyl-5-methyl-1-cyclohexene
24. Indicate which of the following is the ortho structure, the meta structure, and the para structure:

Answer Key

1. a
2. c
3. a
4. a
5. c
6. a
7. d
8. c
9. b
10. c
11. true
12. false
13. true
14. true
15. false
16. false
17. true
18. false
19. true
20. false

1. cycloalkane
2. cycloalkane
3. benzene derivative
4. cycloalkane

1. 1, 2, 3-trimethylpentane
2. 1,1-dichloro-3-methylcyclobutane
3. cycloheptane
4. fluorobenzene

1. para
2. ortho
3. meta
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice

1. Aldehydes all have a
   1. carbonyl group on C-2
   2. carboxyl group on C-2
   3. carbonyl group on C-1
   4. carboxyl group on C-1

2. R-O-R’ is the general formula for
   1. an ester
   2. an ether
   3. an acid
   4. an alcohol

3. A tertiary alcohol has ____ R group(s) attached to the C-OH carbon.
   1. two
   2. three
   3. one
   4. four

4. Amines have a functional group that can contain all the following atoms except
   1. O
   2. H
   3. C
   4. N

5. A class of compounds that are involved in ozone depletion are called
   1. CCFs
   2. CFCs
   3. CFFs
   4. FCCs

6. Glycerol is frequently used as
   1. antifreeze
   2. antiseptic
   3. moisturizer
   4. hand sanitizer

7. Compounds containing the group end their names with
   1. -ol
   2. -al
3. –oic
4. –one

8. Aldehydes and ketones are polar molecules because of the polarized ____ group.
   1. alcohol
   2. carbonyl
   3. carboxyl
   4. amine

9. Propanoic acid is used as
   1. food flavoring
   2. food preservative
   3. paint manufacture
   4. production of adhesives

10. Alcohols interact with water through
    1. strong H bonds
    2. dispersion forces
    3. weak hydrogen bonds
    4. ion-dipole interactions

**True/False**

11. _____ Alcohols can ionize to form OH\(^-\).
12. _____ The carboxylate group is formed when an ester ionizes.
13. _____ Secondary amines have two R groups attached to the N.
14. _____ Carboxylic acids can form H-bonds with water.
15. _____ Ethers interact by weak intermolecular forces.
16. _____ Alcohols bond to each other by dispersion forces.
17. _____ Another name for methanal is formaldehyde.
18. _____ Amines have pleasant odors.
19. _____ Propyl ethanoate smells like pears.
20. _____ Short-chain ketones are not water-soluble.

**Fill in the Blank**

21. Fill in the missing parts of the table below.
22. Name the following compounds:
23. Draw the following structures:
   a. 2-chloro-3-fluorohexane
   b. diethylamine
   c. 4-bromo-2-hexanone
   d. trichloroacetic acid
   e. propyl butanoate

**Answer Key**

1. c
2. b
3. b
4. a  
5. b  
6. c  
7. d  
8. b  
9. b  
10. a  
11. false  
12. false  
13. true  
14. true  
15. true  
16. false  
17. true  
18. false  
19. true  
20. false  

1. propyl hexanoate  
2. 2-propanol  
3. 3-pentanone  
4. 2-bromo-5-iodoheptane  
5. butyl ethyl ether
25.3 Organic Reactions

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice

1. The catalyst for the reaction of methane with chlorine gas is
   1. infrared light
   2. platinum metal
   3. ultraviolet light
   4. iron metal

2. \( \text{Br}_2 \) will react with benzene in the presence of
   1. high temperatures
   2. iron catalyst
   3. high pressures
   4. magnesium catalyst

3. Addition reactions will take place with all of the following except
   1. 2-hexanone
   2. 3-hexyne
   3. propanal
   4. 2-chlorobutane

4. A test for unsaturation uses
   1. \( \text{Br}_2 \)
   2. \( \text{Cl}_2 \)
   3. \( \text{F}_2 \)
   4. \( \text{I}_2 \)

5. The oxidation of a secondary alcohol yields
   1. a ketone
   2. no reaction
   3. an aldehyde
   4. an alkene

6. When heated in the presence of a catalyst, propane will form
   1. propanol
   2. propanone
   3. propene
   4. propanal

7. A dipeptide is formed when the amino group of one amino acid reacts with the ___ group of another amino acid.
   1. carboxyl
2. amino  
3. hydroxyl  
4. carbonyl  

8. The inorganic acid used in the reaction to form an ester is  
   1. HCl  
   2. H₂SO₄  
   3. H₃PO₄  
   4. HBr  

9. Molecules of ______ react to form polymers.  
   1. alkanes  
   2. alkynes  
   3. alkenes  
   4. alcohols  

10. Nylon-66 is formed through the reaction of adipic acid and  
    1. hexanediamine  
    2. hexaneamine  
    3. heaxanetriamine  
    4. aminohexane  

True/False  

11. _____ MgCl₂ is a catalyst for adding alkyl groups to benzene rings.  
12. _____ Water can be added to an alkyne to form a ketone.  
13. _____ Methanol is a product of the reaction between iodomethane and NaOH.  
14. _____ Oxidation of tertiary butyl alcohol yields an aldehyde.  
15. _____ Removal of hydrogen is an oxidation reaction.  
16. _____ In the presence of KOH, an ester will hydrolyze to form a carboxylic acid and an aldehyde.  
17. _____ Monomers of addition polymers contain halogen groups.  
18. _____ A Pt catalyst is required for the hydrogenation of benzene.  
19. _____ PET is formed from a reaction between methylene glycol and terephthalic acid.  
20. _____ Esters are formed by reactions between carboxylic acids and alcohols.  

Fill in the Blank  

21. List the two components that react to form the following esters:  
    1. 2-chloropropyl benzoate  
    2. 3-methylhexyl decanoate  
    3. 4-phenyloctyl propanoate  
22. Complete the following reactions:  
23. Name the starting materials for the following products:  
24. Name the polymer formed by the following monomers:  
    1. CH₂=CH₂  
    2. isoprene  
    3. adipic acid + hexanediamine  
    4. ethylene glycol and terephthalic acid
Answer Key

1. c
2. b
3. d
4. a
5. a
6. c
7. a
8. b
9. c
10. a
11. false
12. false
13. true
14. false
15. true
16. false
17. false
18. true
19. false
20. true

1. benzoic acid and 2-choropropanol
2. decanoic acid and 3-methylhexanol
3. propanoic acid and 4-phenyloctanol

1. propene
2. 1-butene
3. 2-butanol
4. methylbenzene

1. polyethylene
2. polyisoprene
3. nylon-66
4. polyethylene terephthalate (PET)
Chapter Test

Name _____________________ Class ______________________ Date ________________

Multiple Choice

1. Alkynes have the general formula
   1. $C_nH_{2n-2}$
   2. $C_nH_{2n+2}$
   3. $C_nH_{2n}$
   4. $C_nH_{2n-1}$

2. is an example of ________ model for showing organic molecules
   1. ball-and-stick
   2. orbital-filling
   3. structural
   4. space-filling

3. Carbon will form covalent bonds with
   1. Xe
   2. S
   3. Sb
   4. He

4. The bond angles in cyclobutane are
   1. $60^\circ$
   2. $108^\circ$
   3. $109.5^\circ$
   4. $90^\circ$

5. A meta substituent on a benzene ring means that there are ____ attached to the ring.
   1. three groups
   2. four groups
   3. two groups
   4. one group

6. A secondary alcohol has ____ R group(s) attached to the C-OH carbon.
   1. two
   2. three
   3. one
   4. four

7. Esters have a functional group that contains all of the following atoms except
   1. O
   2. H
3. C
4. N

8. Compounds containing the group end their names with
   1. -ol
   2. -al
   3. -oic
   4. -one

9. The oxidation of a tertiary alcohol yields
   1. a ketone
   2. no reaction
   3. an aldehyde
   4. an alkene

10. One of the compounds used in the synthesis of PET is
    1. methylene glycol
    2. polypropylene glycol
    3. ethylene glycol
    4. butylene glycol

**True/False**

11. _____ Alkenes have sp² hybridization.
12. _____ Hepta- is a prefix meaning six.
13. _____ Optical isomers are non-superimposable mirror images of one another.
14. _____ Cyclopentene has two double bonds in the ring.
15. _____ Cyclohexene is more reactive than benzene.
16. _____ Alcohols can form hydrogen bonds with water.
17. _____ Long-chain alcohols are water-soluble.
18. _____ Primary amines have one R group attached to the N.
19. _____ Addition of hydrogen is an oxidation reaction.
20. _____ AlCl₃ is a catalyst for adding alkyl groups to benzene rings.

**Fill in the Blank**

21. Name the following compounds:
22. Draw the following structures:
   a. dichloropropionic acid
   b. ethyl propionate
   c. diethylamine
   d. 2-methylbutanal
23. Write the products of the following reactions:
24. Name the compounds needed to form the following products:
   1. cyclohexane
   2. 3-octanone
   3. pentyl hexanoate
   4. polyethylene
Answer Key

1. a
2. a
3. b
4. d
5. c
6. a
7. d
8. b
9. b
10. c
11. true
12. false
13. true
14. false
15. true
16. true
17. false
18. true
19. false
20. true

1. cycloheptanone
2. 3-chlorobutanoic acid
3. 3-hexanone
4. 4-fluoro-2-pentene

1. benzene + hydrogen gas
2. 3-octanol
3. penatnol + hexanoic acid
4. ethane
Chapter Outline

26.1 Carbohydrates
26.2 Amino Acids and Proteins
26.3 Lipids
26.4 Nucleic Acids
26.5 Biochemistry
26.1 Carbohydrates

Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice

1. Sugars are called carbohydrates because
   1. they contain water molecules
   2. they are hydrated compounds
   3. the generic formula is CH₂O
   4. they are formed from a reaction with H₂O

2. One of the following is not a monosaccharide
   1. glucose
   2. mannose
   3. fructose
   4. galactose

3. In water, the predominant form of glucose is
   1. six carbon straight-chain
   2. six carbon open-chain
   3. six carbon cyclic
   4. six carbon aldehyde

4. A condensation reaction allows glucose and fructose to form a
   1. polysaccharide
   2. disaccharide
   3. extended saccharide
   4. cyclic saccharide

5. Maltose is composed of
   1. glucose + galactose
   2. glucose + lactose
   3. glucose + fructose
   4. glucose + glucose

6. Lactose is found in
   1. table sugar
   2. milk
   3. honey
   4. plants

7. When glucose and galactose combine, the disaccharide formed is
   1. maltose
   2. mannose
3. lactose
4. lactulose

8. The polysaccharide used for storing energy in the body is
   1. glycogen
   2. starch
   3. glycosidase
   4. glucagon

9. Amylopectin is a branched polysaccharide composed of _______ units
   1. fructose
   2. glucose
   3. galactose
   4. lactose

10. Amylose is the _______ form of starch
    1. branched-chain
    2. cyclic-chain
    3. straight-chain
    4. aliphatic

True/False

11. _____ Glucose is found in grapes.
12. _____ The straight-chain form of fructose is an aldehyde.
13. _____ The cyclic form of glucose is a five-membered ring.
14. _____ Fructose + galactose combine to form the disaccharide lactose.
15. _____ The cyclic structures of monosaccharides combine to form disaccharides.
16. _____ Sucrose is often referred to as table sugar.
17. _____ Glucose, fructose, and galactose all have six carbons in their structure.
18. _____ Amylopectin is composed of galactose monomers.
19. _____ Cellulose is a major component of plant fibers.
20. _____ Glycogen is primarily stored in liver and muscle cells.

Fill in the Blank

21. Draw the open-chain structures of glucose and fructose.
22. Draw the cyclic structures of glucose and fructose.
23. Other than the carbonyl group location, what is the major structural difference between glucose and fructose?
24. What is the principal difference between galactose (structure below) and glucose?
25. What is the principal difference between ribose (structure below) and fructose?
26. Indicate which monosaccharide is galactose in the following disaccharide structure.
27. Name this disaccharide:
28. Which of the polysaccharides is the most highly branched?
29. Why is glycogen a good storage form of energy?
30. Amylose is readily digested by humans but cellulose is not. Are there any structural features in the two molecules that might explain this difference?

Answer Key

1. c
26.1. Carbohydrates

2. b
3. c
4. b
5. d
6. b
7. c
8. a
9. b
10. c
11. true
12. false
13. false
14. false
15. true
16. true
17. true
18. false
19. true
20. true
21. Glucose forms a six-membered cyclic structure, while fructose has a five-membered ring.
22. The one –OH group in galactose (see arrow) is oriented up instead of down.
23. Ribose has five carbons and fructose has six.
24. Galactose is the monosaccharide on the left-hand side of the structure.
25. The structure has two glucose monomers and is named maltose.
27. It consists entirely of glucose monomers.
28. The orientations of the –O- bonds. In amylose, the bonds all point below the plane of the ring. In cellulose, alternate O atoms point above the plane of the ring.
Lesson Quiz

Name _____________________ Class ______________________ Date ________________

Multiple Choice

1. All of the following are true about amino acids except
   1. a central carbon is present
   2. the amino group is a secondary amine
   3. the carboxyl group can ionize at basic pH
   4. the R group determines the identity of the specific amino acid

2. R groups can be all of the following except
   1. polar
   2. acidic
   3. basic
   4. halogenated

3. Glu is the abbreviation for the amino acid
   1. glutamic acid
   2. glutamine
   3. glycine
   4. gibberellic acid

4. The R group for tryptophan is
   1. non-polar
   2. cyclic
   3. acidic
   4. methylated

5. The amide bond can also be called the
   1. amine link
   2. peptide chain
   3. peptide bond
   4. amino bond

6. The amino acid chain glutamic acid-tyrosine-cysteine-threonine can be abbreviated as
   1. glu-tyr-cys-thr
   2. thr-cys-tyr-glu
   3. gln-tyr-cys-thr
   4. thr-cys-tyr-gln

7. The beta sheet is an example of
   1. tertiary structure
   2. primary structure
26.2. Amino Acids and Proteins

3. secondary structure
4. quaternary structure

8. The substrate is converted to product at the _______.
   1. reactive site
   2. active site
   3. reaction site
   4. activity site

9. The hemoglobin molecule illustrates _______ structure
   1. tertiary structure
   2. primary structure
   3. secondary structure
   4. quaternary structure

10. Magnesium can serve as an enzyme
    1. coreactant
    2. cofactor
    3. coproduct
    4. cosubstrate

True/False

11. _____ There are twenty amino acids associated with living organisms.
12. _____ All amino acids have R-groups.
13. _____ The side chain for threonine contains a carboxyl group.
14. _____ Peptide bond formation occurs between two carboxyl groups on different amino acids.
15. _____ Trp-Thr-Tyr indicates the amino acid sequence tryptophan-threonine-tyrosine.
16. _____ Hemoglobin contains an iron atom in the center of each subunit.
17. _____ An inhibitor enhances substrate binding to the active site.
18. _____ Substrate is converted to product at the active site.
19. _____ Cyanide is an inhibitor of the enzyme cytochrome c oxidase.
20. _____ Cofactors interfere with the activity of an enzyme.

Fill in the Blank

21. Write the name of the amino acid represented by the following abbreviations:
    1. ala
    2. trp
    3. ile
    4. leu

22. Write the abbreviation for each of the following amino acids:
    1. asparagine
    2. cysteine
    3. lysine
    4. arginine

23. Draw the R-groups for alanine, threonine, and lysine.
24. Indicate how many ways alanine and aspartic acid can form amide bonds.
25. Describe the structure of hemoglobin.
26. How is an alpha helix represented in the hemoglobin structure?
27. How would a beta sheet be represented?
28. In the following reaction, name the substrate and the product:

29. When writing enzyme reactions, where does the name of the enzyme go?

30. Acetylcholine (AC) is involved in transmission of nerve impulses. After an impulse is triggered by release of AC, the compound is quickly inactivated by an enzyme by breaking the amide bond. Carbamate serves as an inhibitor of the enzyme. Why is carbamate an effective inhibitor (hint: look at the two structures)?

---

**Answer Key**

1. b
2. d
3. a
4. b
5. c
6. a
7. c
8. b
9. d
10. b
11. true
12. true
13. false
14. false
15. true
16. true
17. false
18. true
19. true
20. false

1. alanine
2. tryptophan
3. isoleucine
4. leucine

1. asn
2. cys
3. lys
4. arg

1. amine of alanine to carboxyl of aspartic acid
2. carboxyl of alanine to amine of aspartic acid
3. amine of alanine to R-group carboxyl of aspartic acid

21. Hemoglobin is composed of two alpha subunits and two beta subunits. Each subunit has a porphyrin ring that holds an iron atom in its center.

22. The alpha helix is represented by a spiral:

23. The beta sheet is represented by a ribbon with an arrow:

24. Asparagine is the substrate and aspartic acid is the product.

25. Over the arrow

26. The left-hand sides of both molecules contain oxygen structures that are identical. In acetylcholine, the amide bond is broken in the reaction. The inhibitor does not have an amide bond, so it will not break down. The
inhibitor structure is very similar to that of the substrate, so it will fit in the active site and block substrate from reacting.
Lesson Quiz

1. Triglycerides are classified as
   1. carbohydrates
   2. lipids
   3. proteins
   4. phospholipids

2. A major source for canola oil is
   1. animal fat
   2. synthesis in the lab
   3. plants
   4. petroleum products

3. Oleic acid is
   1. saturated
   2. monounsaturated
   3. diunsaturated
   4. polyunsaturated

4. Foods containing higher amounts of unsaturated fats include
   1. butter
   2. cheese
   3. lard
   4. nuts

5. The highest amount of unsaturated fat is found in
   1. almond
   2. olive oil
   3. canola oil
   4. mayonnaise

6. In water, phospholipids spontaneously form a
   1. lipid monolayer
   2. lipid bilayer
   3. lipid hydrolayer
   4. lipid layer

7. The hydrophilic portion of a phospholipid is
   1. unsaturated fatty acid
   2. phosphate
3. phosphate plus ionic group  
4. saturated fatty acid

8. Bilayers allow materials to pass through because of
   1. proteins on the interior of the bilayer
   2. proteins on the exterior of the bilayer
   3. protein channels
   4. carbohydrates on the bilayer surface

9. One of the following esters would not be considered a wax
   1. cetyl palmitate
   2. cetyl acetate
   3. cetyl linoleate
   4. cetyl oleate

10. Waxes have all of the following characteristics except
    1. soft solid
    2. low melting point
    3. water-soluble
    4. fatty acid ester

True/False

11. _____ Fatty acids do not dissolve in water.
12. _____ Canola oil is a vegetable oil.
13. _____ All triglycerides contain unsaturated fatty acids.
14. _____ The majority of fats in mayonnaise are saturated fats.
15. _____ Phospholipids contain three fatty acids.
16. _____ The phosphate portion of a phospholipid is polar.
17. _____ Phospholipids are minor components of cell membranes.
18. _____ Carbohydrate groups are attached to the outside of cell membranes.
19. _____ Waxes help plants absorb water.
20. _____ Waxes are found on skin and hair.

Fill in the Blank

21. Name the acid found in this triglyceride:
22. Name the saturated and monounsaturated fatty acids in the structure below:
23. Draw the structure of the triglyceride containing three molecules of heptanoic acid.
24. Are the double bonds in the structures below cis or trans?
25. What structural properties of the phospholipids shown below make the head hydrophilic?
26. Lauric acid is a twelve-carbon saturated fatty acid. Draw the structure of the ethyl ester of lauric acid. Would this ester be considered a wax?
27. Why are unsaturated fats considered to be healthier than saturated fats?
28. Protein channels transport materials in and out of cells. Why do these channels need to be protein in structure?
29. What properties of waxes make them water-repellant?
30. How could carbohydrates help identify types of cells?

Answer Key

1. b
2. c
3. b
4. d
5. c
6. b
7. c
8. c
9. b
10. c
11. true
12. true
13. false
14. false
15. false
16. true
17. false
18. true
19. false
20. true
21. decanoic acid
22. saturated: palmitic; unsaturated: oleic
   1. trans
   2. cis
23. The O\textsuperscript{−} and the N can both interact with water molecules.
24. It is not a wax because both the acid and alcohol chains are too short.
25. The unsaturated fats contain fewer calories.
26. The molecules and ions that move in and out of cells are water-soluble and need a hydrophilic environment in which to move. Proteins can provide a channel that allows water to move.
27. The long-chain fatty acid and alcohol are both very hydrophobic.
28. Different carbohydrate structures (either different monoaccahrides or olysaccharides) could represent different types of cells.
26.4 Nucleic Acids

Lesson Quiz

Name _____________________  Class ______________________  Date ________________

Multiple Choice

1. A major responsibility of DNA is to
   1. form proteins
   2. direct how cells make proteins
   3. provide flexibility to the cell
   4. direct cell division

2. RNA has the responsibility to
   1. organize cell contents
   2. store energy for the cell
   3. facilitate protein synthesis
   4. store information

3. The carbohydrates in both RNA and DNA contain
   1. five carbons
   2. six carbons
   3. unsaturation
   4. straight-chain structure

4. The nitrogenous base in both RNA and DNA attaches covalently to
   1. phosphate group
   2. another base
   3. carbohydrate
   4. amino groups

5. The covalent bonds that link the individual bases together are between
   1. base and base
   2. phosphate and carbohydrate
   3. phosphate and phosphate
   4. carbohydrate and carbohydrate

6. One of the hydrogen bonds between thymine and adenine is between
   1. O on thymine and NH on adenine
   2. O on thymine and N on adenine
   3. N on adenine and O on thymine
   4. O on adenine and N on thymine

7. All of the following are stop codes except
   1. TAA
   2. TGA
3. TAG
4. TAT

8. The sequence AAA-ATA-TTG represents the amino acid sequence
   1. lys-ile-leu
   2. lys-ile-phe
   3. ile-lys-leu
   4. leu-ile-lys

9. The amino acid sequence cys-leu-ala is indicated by the following
   1. TGT-CTG-AGG
   2. TGC-CTA-GCG
   3. TGT-CTT-AAC
   4. TGC-ATC-GCG

10. All of the following are true about a gene except
    1. it is a segment of DNA
    2. it codes for a specific polypeptide chain
    3. it provides structure for the nucleus
    4. it gives information about what the organism needs in order to function

True/False

11. _____ DNA and RNA are found in the cell nucleus.
12. _____ The carbohydrate in DNA is ribose.
13. _____ Cytosine has two sites for H-bonding.
14. _____ In RNA, thymine is replaced by uracil.
15. _____ Adenine and guanine are two bases found in DNA.
16. _____ DNA has a single-stranded structure.
17. _____ The genetic code is a four-letter code.
18. _____ All amino acids have more than one base code.
19. _____ Stop codes direct the finish of a protein.
20. _____ In a DNA chain, guanine will form H-bonds with cytosine and adenine.

Fill in the Blank

21. A ______ is a molecule that contains a five-carbon ______, a phosphate group, and a ______-containing base called a nucleobase.
22. Draw the structure of the nucleotide containing cytosine as the base.
23. The structure for uracil is given below. Draw the structure for the uracil-adenine hydrogen bonding.
24. Write the amino acid sequences that would be generated by the following base sequences:
   a. TTC-GGG-CAC-GTG
   b. AAA-ACA-TAC-GCA
25. Write the triplet codes that would direct the incorporation of serine into a protein chain.
26. A ______ is a segment of DNA that carries a ______ for making a specific ________ chain.
27. What is a triplet?
28. Write the DNA code and the amino acid abbreviation for an amino acid that has only one triplet coding for it.
29. Where on the five-membered carbohydrate ring does the phosphate attach?
30. How many different two-base combinations are there for the synthesis of asn-glu?
26.4. Nucleic Acids

Answer Key

1. b
2. c
3. a
4. c
5. b
6. a
7. d
8. a
9. d
10. c
11. true
12. false
13. false
14. true
15. true
16. false
17. false
18. false
19. true
20. false
21. nucleotide, sugar, nitrogen-
    1. phe-gly-his-val
    2. lys-thr-tyr-ala
22. There are six different sequences: AGT, AGC, TCA, TCG, TCT, TCC
23. gene, code, polypeptide
24. A sequence of three letters that designates a base sequence coding for a specific amino acid.
25. ATG, methionine; or TGG, tryptophan
26. On a free –OH group
27. Four –AAT-GAA, AAT-GAG, AAC-GAA, AAC-GAG
26.5 Biochemistry

Chapter Test

Name _____________________ Class ______________________ Date ________________

Multiple Choice

1. All of the following are true about glycogen except
   1. glycogen is composed of glucose subunits
   2. glycogen is stored in liver and muscle cells
   3. glycogen is a trisaccharide
   4. glycogen provides quick energy for the body

2. Sucrose is composed of
   1. glucose + galactose
   2. glucose + lactose
   3. glucose + fructose
   4. glucose + glucose

3. Amylopectin is the _____ form of starch
   1. branched-chain
   2. cyclic-chain
   3. straight-chain
   4. aliphatic

4. Gln is the abbreviation for the amino acid
   1. glutamic acid
   2. glutamine
   3. glycine
   4. gibberellic acid

5. _____ in red blood cells is an important body protein
   1. hemoglobin
   2. hemaglobin
   3. hemoglobin
   4. hemogloban

6. An enzyme is a biochemical
   1. catalyst
   2. cofactor
   3. reactant
   4. product

7. Triglycerides are composed of
   1. fatty acids
   2. fatty acids + glycerol
3. fatty acids + carbohydrates
4. glycerol

8. One of the following would be considered a wax
   1. ethyl palmitate
   2. cetyl butyrate
   3. propyl linoleate
   4. cetyl oleate

9. All of the following are DNA bases except
   1. adenine
   2. cytosine
   3. uracil
   4. guanine

10. DNA and RNA differ in all of the following except
    1. carbohydrate structure
    2. helical structure
    3. base structure
    4. phosphate bridge

True/False

11. _____ The cyclic form of fructose is a five-membered ring.
12. _____ Glucose is also referred to as grape sugar.
13. _____ The straight-chain form of glucose is a ketone.
14. _____ The side-chain for lysine contains a carboxyl group.
15. _____ An inhibitor blocks access by substrate to the active site.
16. _____ Peptide bond formation takes place with the R-groups of amino acids.
17. _____ Butter is a fat derived from animals.
18. _____ Triglycerides are used for quick energy by the body.
19. _____ The carbohydrate in RNA is ribose.
20. _____ DNA has a double-helical structure.

Fill in the Blank

21. Define the following terms:
   a. carbohydrate
   b. polysaccharide
   c. peptide bond
   d. lipid
   e. nucleotide

22. The open-chain structure for galactose is shown below. Draw the cyclic structure:

23. Which of the following compounds fit the definition of an amino acid? Which one would you expect to be used to form a protein? Explain your answer.

24. The following is a part of a silk protein structure. Which level of structure is illustrated here?

25. What levels of structure can be observed in the following protein?

26. For the following enzyme reaction, identify:
   a. the substrate (describe the compound)
   b. the product (describe the compound)
   c. the enzyme
27. Name the fatty acid that is in this triglyceride.
28. Would octyl decanoate be classified as a wax? Explain your answer.
29. Write the amino acid sequence coded for by the nucleic acid sequence: TGG-CTC-ATG-ACA.
30. Given the structures for adenine and uracil, what would a base pair look like if these two compounds formed H-bonds? Assume that adenine connects with the rest of the chain at the N in the five-membered ring that does not have a double bond. Uracil will connect to the rest of the chain at the N at the bottom of the structure (not the one between the two C=O groups).

Answer Key

1. c
2. c
3. a
4. b
5. c
6. a
7. b
8. d
9. c
10. d
11. true
12. true
13. false
14. false
15. true
16. false
17. true
18. false
19. true
20. true

1. monomers and polymers of aldehydes and ketones that have multiple hydroxyl groups attached.
2. a complex carbohydrate polymer formed from the linkage of many monosaccharide monomers.
3. the amide bond that occurs between the amino nitrogen of one amino acid and the carboxyl carbon of another amino acid.
4. a class of water-insoluble compounds that includes oils, fats, and waxes.
5. a molecule that contains a five-carbon sugar, a phosphate group, and a nitrogen-containing base called a nucleobase.

21. b and c would fit the broader definition, but c is the only structure that would be used in a protein chain, since it is the only one that has the carboxyl group and the amino group both attached to the same carbon. A might be considered, even though it has a secondary amine group.
22. primary structure
23. overall, tertiary structure. Elements of secondary structure can be seen in the alpha helices and the beta sheets.
24. The carboxyl compound on the left is the substrate and the ketone on the right is the product. The enzyme is tyrosinase.
25. Octanoic acid.
26. No, both the acid and alcohol chains are too short.
27. trp-leu-met-thr