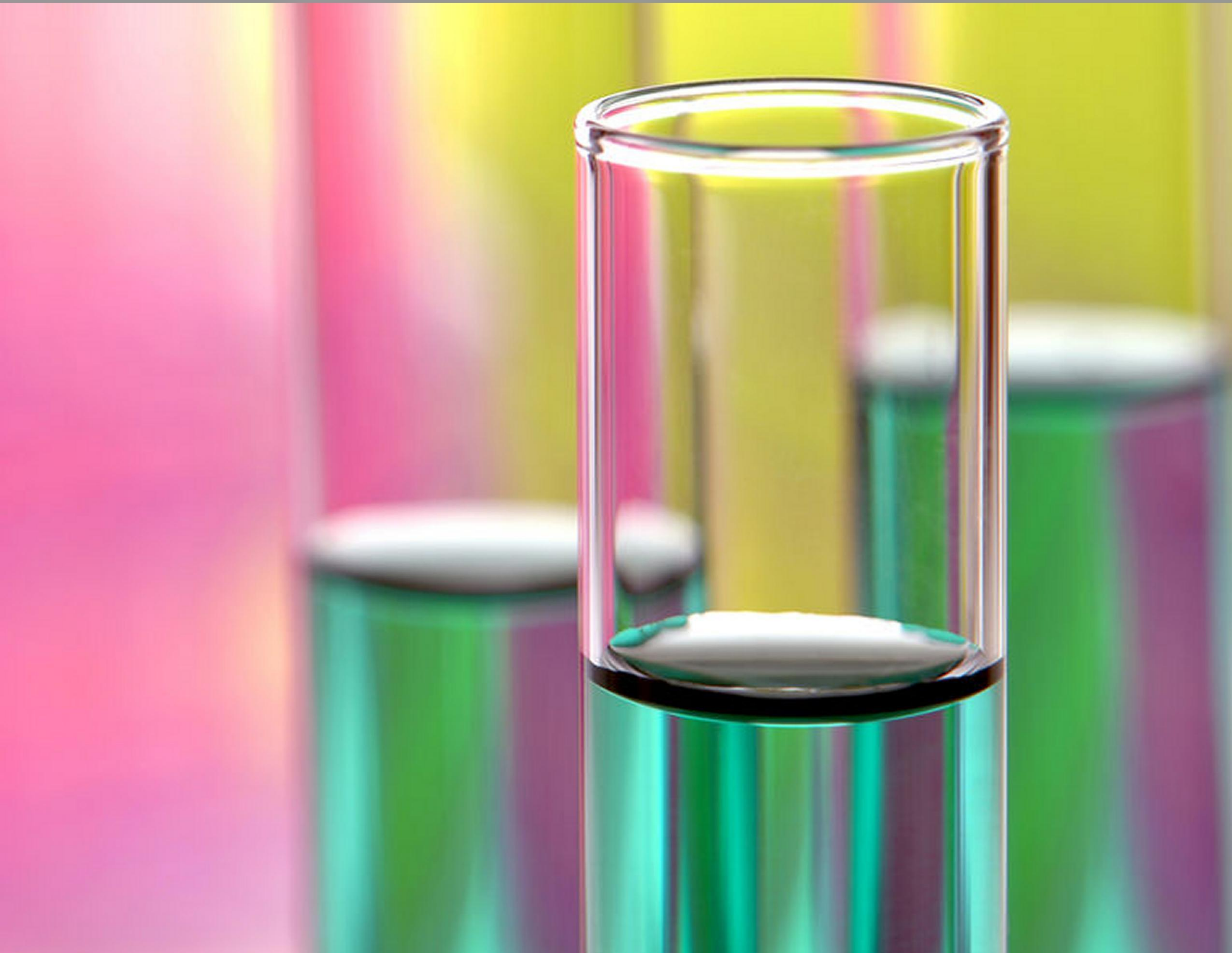
CK-12 Chemistry - Intermediate Quizzes and Tests



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Donald Calbreath, Ph.D.

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AUTHOR
Donald Calbreath, Ph.D.

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CHAPTER 1 Introduction to Chemistry Assessments

Chapter Outline

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?

1.2 The Scientific Method

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

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.

1.3 Introduction to Chemistry

Chapter Test

Name___________________ Class______________ Date________

Multiple Choice:

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 that 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

   c. copper cables
   d. polymers

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.

CHAPTER 2 Matter and Change Assessments

Chapter Outline

2.1 Properties of Matter

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/cm^3
   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

c. being stretched
d. oxidation

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 __________

b. 104.7°C __________
c. 176.9°C __________

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. H_2SO_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 ($C_{12}H_{22}O_{11}$) 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:

Fraction	Boiling Point (0C)	Use
butane	below 30	natural gas
gasoline	100-150	car and truck fuel
diesel oil	200-300	fuel for trucks and trains
greases and waxes	400-500	lubricants

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

2.3 Changes in Matter

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

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:
22. $2H_2O \rightarrow 2H_2 + O_2$
   a. One reactant in this process is _______.
   b. One product in this process is _______.
23. The color in the test tube changes from _______ to ______ when mercuric oxide is heated.
24. 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 _____________
25. 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?

2.4 Matter and Change

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/cm^3
   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 $NaHCO_3$
   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?

CHAPTER 3 Measurements Assessments

Chapter Outline

3.1 The International System of Units

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. 1.5×10^3 *gigabytes*
   b. 1.5×10^2 *kilobytes*
   c. 1.5×10^{-2} *gigabtyes*
   d. 1.5×10^6 *kilobytes*
4. The prefix hecto has an exponential factor of
   a. 1×10^2
   b. 1×10^{-2}
   c. 1×10^4
   d. 1×10^{-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

c. 173.15
d. 473.15

8. The number 0.0015 can be represented by

a. 1.5×10^3
b. 1.5×10^2
c. 1.5×10^{-3}
d. 1.5×10^{-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. 67.3×10^2 _____
c. 0.0045 _____
d. 8.2×10^{-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)?

3.2 Unit Conversions

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. $length \times mass$
   b. $width \times volume$
   c. $length \times volume$
   d. $length \times width$
5. The units for speed are
   a. m/s
   b. kg/s
   c. m/L
   d. m^2/s
6. Density is
   a. $mass \times volume$
   b. an intensive property
   c. $\frac{volume}{mass}$
   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

c. mass of material
d. density of material

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/m^3.

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: $20\ m \times 10\ m \times 5\ m$. 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?

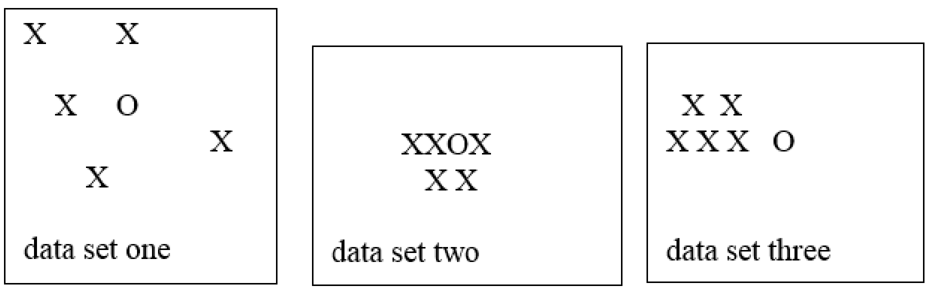
3.3 Uncertainty in Measurements

Lesson Quiz

Name____________________ Class_______________ Date_________

Multiple Choice:

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:

TABLE 3.1:

Experiment	Accepted Value	Experimental Value	Percent Error
a.	14.85	12.99	
b.	14.85	16.32	
c.	14.85		+2.7

20. Fill in the following table:

TABLE 3.2:

Number	How Many Significant Figures?
a. 7.2×10^{-4}	
b. 33.709	
c. 1408	
d. 2.69×10^{3}	

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?

TABLE 3.3:

Experiment	Accepted Value	Experimental Value	Percent Error
a.	14.85	12.99	12.5%
b.	14.85	16.32	9.9%
c.	14.85	15.25	+2.7

TABLE 3.4:

Number	How Many Significant Figures?
a. 7.2×10^{-4}	two
b. 33.709	five
c. 1408	four
d. 2.69×10^{3}	three

3.4 Measurements

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:

Analysis	Technician One	Technician Two	Technician Three
1	12.98	13.98	11.45
2	17.62	14.22	11.63
3	10.51	13.87	12.09
4	7.35	14.15	11.84

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

c. 5000
d. 0.5

6. Density is
   a. $\frac{mass}{volume}$
   b. $\frac{volume}{mass}$
   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. 1×10^{-9}
   b. 1×10^{-6}
   c. 1×10^{-3}
   d. 1×10^{-12}
9. To convert from kilograms to grams, multiply the kilogram value by
   a. 10
   b. 1×10^{2}
   c. 1×10^{2}
   d. 1×10^{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 μL.
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 μL.
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 $3.0 \times 10^8\ m/s$. 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.

CHAPTER 4 Atomic Structure Assessments

Chapter Outline

4.1 Atoms

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

c. additive proportions
d. conservation of mass

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 that 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 you answer.

4.2 The Nuclear Model of the Atom

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

c. cathane rays
d. chain

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 neuron 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?

4.3 Isotopes and Atomic Mass

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. $_{28}$Si
   d. $\overset{28}{Si}$
5. How many neutrons are in the nucleus of $^{96}_{42}Mo$?
   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

c. atomic mass unit
d. atomic weight unit

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. neuron 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 $_{82}Pb$ indicates an element with an atomic mass of 82.
15. _____ Ge-73 is another way of showing atomic mass.
16. _____ The isotope $^{101}_{44}Ru$ 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 4.1:

Element	Symbol	Protons	Neutrons	Electrons	Atomic Number	Atomic Mass
Indium	In			49		114
Yttrium	Y		50		39	
Sulfur	S				16	32
Rubidium	Rb	37				85
Argon	Ar		22	18		

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?

TABLE 4.2:

Element	Symbol	Protons	Neutrons	Electrons	Atomic Number	Atomic Mass
Indium	In	49	65	49	49	114
Yttrium	Y	39	50	39	39	89
Sulfur	S	16	16	16	16	32
Rubidium	Rb	37	48	37	37	85
Argon	Ar	18	22	18	18	40

4.4 Atomic Structure

Chapter Test

Name___________________ **Class**______________ **Date**________

Multiple Choice:

1. Which of the following reactions illustrates the law of conservation of mass?
   a. 43.2 g A + 17.9 g B → 30.1 g C + 31.0 g D
   b. 68.2 g A → 27.1 g B + 43.5 g C
   c. 12.3 g A + 9.6 g B → 23.9 g C
   d. 18.7 g A + 22.4 g B → 26.3 g C + 21.6 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 $^{108}_{47}Ag$ 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:

Element	Symbol	Protons	Neutrons	Electrons	Atomic Number	Atomic Mass
cesium	$^{129}_{55}Cs$					
iridium	Ir	77				192
arsenic	As		41			74
palladium	Pd			46		106

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

carbons and 6 hydrogens. Carbon has an atomic mass of 12 and H has an atomic mass of 1. What is the mass ratio of H to carbon in the two compounds?

27. 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.

TABLE 4.4:

Element	Symbol	Protons	Neutrons	Electrons	Atomic Number	Atomic Mass
cesium	${}^{129}_{55}Cs$	55	74	55	55	129
iridium	Ir	77	115	77	77	192
arsenic	As	33	41	33	33	74
palladium	Pd	46	60	46	46	106

CHAPTER 5 Electrons in Atoms Assessments

Chapter Outline

5.1 Light

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 6×10^{-10} m?
31. What is the frequency of light that has an energy of 3.71×10^{-17} J?

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?

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. μg

b. mg
c. g
d. kg

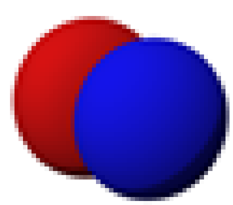
8. An increase in the quantum number n means ________ in the distance of the electron from the nucleus.
   a. an increase
   b. no change
   c. a decrease
   d. a variable response
9. If the principal quantum number $n = 4$, the angular quantum number can have values for zero to
   a. 2
   b. 4
   c. 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 $2.05 \times 10^7\ kg$ travelling at a rate of $7.85 \times 10^3\ m/s$.

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?

5.3 Electron Arrangement in Atoms

Lesson Quiz

Name____________________ Class_______________ Date_________

Multiple Choice:

1. The iodine atom has the electronic configuration $[Kr]4d^{10}5s^{2}5p^{5}$. 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 $_{12}Mg$ 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 $_{14}Si$, the noble gas to use as a shorthand symbol is
   a. $_{86}Rn$
   b. $_{2}He$
   c. $_{10}Ne$
   d. $_{18}Ar$
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

   b. Pauli exclusion principle
   c. Aufbau principle
   d. Scrödinger 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. _____ $_5$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 ($1s^2 2s^2$) 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. $_{19}$K
    b. $_{21}$Sc
    c. $_{15}$P
    d. $_{18}$Ar
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]$4d^4 5s$

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?

5.4 Electrons in Atoms

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]$3d^5 4s$. How many valence electrons does chromium have?
   a. six
   b. one
   c. five
   d. two
5. Wavelength is designated by the symbol
   a. ν
   b. δ
   c. η
   d. λ
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

b. three
c. two
d. four

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 $_{24}Cr$, the noble gas to use as a shorthand symbol is
   a. $_{36}Kr$
   b. $_{10}Ne$
   c. $_{18}Ar$
   d. $_{54}Xe$

True/False:

11. _____ Unpaired electrons may have different spins.
12. _____ $_5B$ ($1s^22s^22p$) 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. _____ $_{18}Ar$ 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. $_{22}Ti$
   b. $_{19}K$
   c. $_{13}Al$
24. Use arrows to indicate the orbitals for the following –ignore the noble gas component
   a. vanadium: $[Ar]3d^34s^2$
   b. germanium (valence elecrons only): $[Ar]3d^{10}4s^24p^2$
   c. polonium: $[Hg]6p^4$
25. Calculate the wavelength of a bumblebee with a mass of 1×10^{-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?

CHAPTER 6 The Periodic Table Assessments

Chapter Outline

6.1 History of the Periodic 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.

TABLE 6.1: Class Period(s) (60 min)

Test	Metal	Nonmetal	Metalloid
heat conductivity	good	very poor	some
electrical conductivity	good	very poor	some
malleability	very malleable	little malleability	some
ductility	very ductile	little ductility	some

6.2 Electron Configuration and the Periodic 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 $[Kr]4d^5 5s^2$ appears in period
   1. seven
   2. five
   3. three
   4. four
4. An element with the electron configuration $[Ar]3d^5 4s^2$ 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 $1s^2$ 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 $[Ar]3d^14s^2$, 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 6.2: Class Period(s) (60 min)

Configuration	Period	Group	Block	Valence Electrons
$[Xe]4f^{14}5d^56s^2$				
$[Ne]3s^23p$				
$1s^22s^22p^3$				

Short Answer:

30. Why is helium a group 18 element?

TABLE 6.3: Class Period(s) (60 min)

Configuration	Period	Group	Block	Valence Electrons
$[Xe]4f^{14}5d^{5}6s^{2}$	six	six	d	two
$[Ne]3s^{2}3p$	three	thirteen	p	three
$1s^{2}2s^{2}2p^{3}$	two	fifteen	p	five

6.3 Periodic Trends

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. ns^2
   2. np^4
   3. ns^2p^3
   4. ns^2p^4
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 patters are observed with regard to ionization energies (IE):

TABLE 6.4: Class Period(s) (60 min)

Element	IE1	IE2	IE3	IE4
Li	520	7300	——	——
B	801	2430	3660	25,000
Mg	738	1450	7730	——

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

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. $_{40}Zr$ 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. ns^2np^3
   2. $nsnp^4$
   3. ns^2np^2
   4. $nd^{10}s^2p^2$
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. ______ $_{70}Yb$ 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 ns^2np^6 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 $[Xe]4f^{14}5d^9 6s$ represents the element _____.
28. Write the electron configuration for $_{15}P$

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 $_{13}Al$ is 578 kJ/mol. Offer a reasonable explanation for this observation.

CHAPTER 7 Chemical Nomenclature Assessments

Chapter Outline

7.1 Ionic Compounds

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. O^{2-}
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 As^{3-} ion is
    1. arsenic
    2. arsenous
    3. arsenide
    4. arsenate
7. The correct charge for the chromium(III) ion is
    1. Cr^{+}
    2. Cr^{2+}

   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. $Hg_2{}^{2+}$ is a
   1. polyatomic cation
   2. polyatomic compound
   3. monatomic cation
   4. trivalent cation

True/False:

11. _____ Na_2Cl_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_2S
   2. Fe_2O_3
   3. $CuCl_2$
   4. AuBr
   5. MnO
   6. Ag_3SO_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?

7.2 Molecular Compounds

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. CH_3Cl
    3. $Na_2Cr_2O_7$
    4. $CH_3CH_2CH_3$
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. S_2O
    3. SO_2
    4. S_2O_2
7. The formula P_2O_5 represents the compound
    1. pentoxide diphosphate
    2. diphosphorus pentoxide

3. phosphorus pentaoxide
4. diphosphopentoxide

8. The formula BH_3 represents the compound

1. boron trihydride
2. beryllium trihydride
3. boron trihydrate
4. boron trihydogen

9. Tellurium trioxide is represented by the following formula

1. Te_2O_3
2. TeO_3
3. Te_3O_2
4. TeO_4

10. Oxygen difluoride has the formula

1. OF_3
2. O_3F
3. O_3F_3
4. OF_2

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. S_2Cl_2
2. NO_2
3. CCl_4
4. IBr_4
5. BrF_5
6. SiO_3

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?

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_4OH 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. _____ CH_3Cl 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. _____ H_2SO_4 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. HNO_2
    3. $Ba(OH)_2$
    4. RbOH
    5. H_3PO_3
    6. H_2CrO_4
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?

7.4 Chemical Nomenclature

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_3F
    3. O_3F_3
    4. OF_2
7. Selenium dioxide has the following formula
    1. SeO_3
    2. Se_3O_2

3. SeO_2
4. Se_2O_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_3P_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_2SO_3
 3. H_3PO_4
 4. $MnCl_2$
 5. AuI_3
 6. $Ba(NO_3)_2$
 7. H_2SO_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 he exact name and formula for a compound?

CHAPTER 8 Ionic and Metallic Bonding Assessments

Chapter Outline

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. $:\ddot{\underset{\cdot\cdot}{\mathrm{X}}}:$

   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?

8.2 Ionic Bonds and Ionic Compounds

Lesson Quiz

Name _____________________ **Class** ______________________ **Date** ________________

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. $:\ddot{\mathrm{X}}:$

    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 $Al_2(SO_4)_3$ 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. _____

$$: \mathrm{X} :$$

is the period 4 electron dot formula for germanium.

Fill in the Blank:

21. LiCl and $PbCl_2$ 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 $CoCl_2$ have when dissolved in water?
27. ________ charged particles attract one another.
28. An _____ compound is an electrically ______ compound consisting of _____ and _____ ions.

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. polyhedric
   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

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 ______.

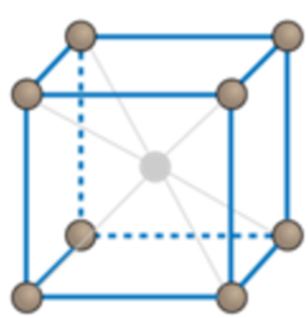
1. Al
2. Cd
3. Cr
4. Cu

True/False:

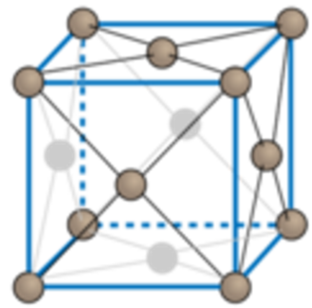
11. _____ The major constituent of steel is iron.
12. _____ Brass is an alloy of copper and tin.
13. _____ Close packing provides room for the electrons to flow easily.
14. _____ Brass is used in the manufacture of trumpets.
15. _____ Se shows a high degree of metallic bonding.
16. _____ The hexagonal 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:

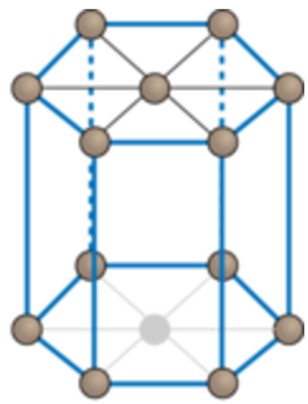
21. An alloy is a mixture of _____ or more elements, at least one of which is a ______.
22. Name the crystal structures:



23. **A**



24. **B**



25. **C**

26. A: ________________________________ B: ________________________________ C: _____________-____________________

27. The model of metallic bonding is called the ___________ model.
28. Alloys are commonly used in _________ objects because the properties of the mixtures are often more _____-__ than the pure metal.
29. 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.

8.4 Ionic and Metallic Bonding

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 $[Ne]3s^2 3p^6$
    1. Na^+
    2. S^{2-}
    3. Mg^{2+}
    4. P^{2-}
3. The formula unit for aluminum sulfate is
    1. $AlSO_4$
    2. $Al_3(SO_4)_2$
    3. $Al(SO_4)_2$
    4. $Al_2(SO_4)_3$
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. $\cdot\mathrm{X}:$

    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^+

    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. _____

$$:\ddot{\underset{\cdot\cdot}{X}}:$$

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_2O_3 is the principal component of rust.

Fill in the Blank:

21. A pseudo noble gas electron configuration contains ___ electrons.
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.

CHAPTER 9

Covalent Bonding Assessments

Chapter Outline

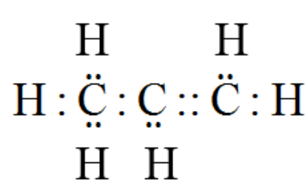
9.1 Lewis Electron Dot Structures

Lesson Quiz

Name____________________ **Class**_______________ **Date** ________

Multiple Choice:

1. One of the following is not a polyatomic ion
    1. SO_2
    2. SO_{3-}
    3. SO_{42-}
    4. HSO_{3-}
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
    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.

1. NO_2
2. PCl_5
3. BH_3
4. $BeCl_2$

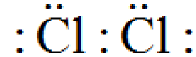
7. The amount of energy needed to break a halogen-halogen bond is
    1. fairly small
    2. high
    3. very high
    4. moderate
8. Which is the proper way to represent the Lewis structure for a polyatomic ion?
    1. $[L^+]$
    2. L^+
    3. $[L]^+$
    4. $L[^+]$
9. In the CO molecule, the coordinate 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

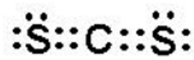
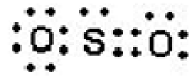
True/False:

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 F_2 molecule contains two shared pairs of electrons.

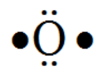
Fill in the Blank:

21. Draw the Lewis structures for
    1. OH^-
    2. NH_3
    3. Cl-Br
22. Draw the structural formulas for

1. 

2. 
3. 

23. Which exception to the octet rule is shown by BH_3. Explain your answer
24. The two identical bonds in O_3 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



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.

9.2 Molecular Geometry

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 AB_3 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 AB_2E 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. SF_5
    2. ClF_3
    3. SF_4
    4. Cl_2F_5
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

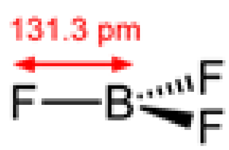
8. The triiodide ion has a _______ geometry
   1. octahedral
   2. linear
   3. square pyramidal
   4. T-shaped
9. The H-O-H bond angle is
   1. 109.5°
   2. 107°
   3. 104.5°
   4. 103.2°
10. A characteristic of the ClF_3 molecule is
    1. two lone pair sets of electrons
    2. one lone pair set of 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 (SF_6) 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 BF_3 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?

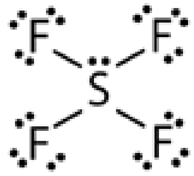


22.

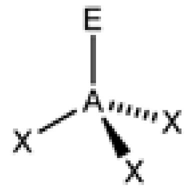
23. Why does the ozone molecule take on a bent configuration?

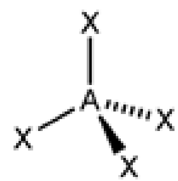
24.

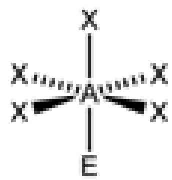
25. The molecular geometry of ammonia is ________.
26. Why does the sulfur tetrafluoride molecule have a distorted tetrahedron configuration?

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)

   1. 

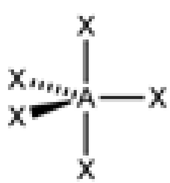
   2. 

   3. 

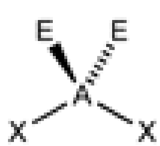
31. Match the structure with the compound

   1. PCl_5
   2. SF_6
   3. H_2O

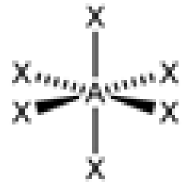
32. **Image 1**

33. 

34. **Image 2**

35. 

36. **Image 3**

37. 

9.3 Polarity and Intermolecular Forces

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. Δ
   2. λ
   3. δ
   4. α
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?

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. sp^2
   3. sp^3
   4. sp^4
6. he electron domain geometry for BF_3 leads to _____ hybridization.
   1. sp
   2. sp^2
   3. sp^3
   4. sp^4
7. The electron domain geometry of sp^2 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. BeH_2
   2. NH_3
   3. CO_2
   4. C_2H_2

9. Octahedral electron domain geometry is seen in _____ hybridization.

   1. sp^2
   2. sp^3d
   3. sp^3d^2
   4. sp^3

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 sp^2, 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:

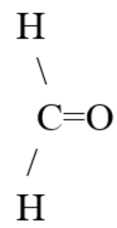
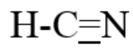
    1. hybridization
    2. valence bond theory
    3. hybrid orbitals
    4. sigma bond
    5. pi bond

22. Describe the bonding that takes place in each of the following diatomic molecules

    1. CS_2
    2. I_2
    3. H_2S

23. Indicate the hybridization of the central carbon atom in each of the following:

    1. CH_4

2. 
3. 

24. How can SF_6 involve d orbitals when there are no d electrons in either atom?

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_2H_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

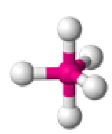
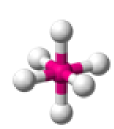
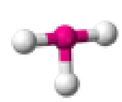
   3. helium
   4. iodine

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. $CHBr_3$ displays ______ hybridization
   1. sp
   2. sp^2
   3. sp^3
   4. sp^4
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 NO_2. Explain your answer.
22. Draw the Lewis structure for
    1. O_2
    2. HBr
    3. F_2
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

CHAPTER 10 The Mole Assessments

Chapter Outline

10.1 The Mole Concept

Lesson Quiz

Name____________________ **Class**_______________ **Date**_________

Multiple Choice:

1. The molecule $CH_2CHCH_3CH_2CHF(CH_3)_2CCH_3$ 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 SiO_3
    4. a molecule of SiO_2
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 O_2 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. on mole of representative particles of a substance
    3. 6.02×10^{20} 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 $CH_2CHCH_3CH_2CHF(CH_3)_2CCH_3$ 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. 1.7×10^{14} atoms of Pt
    2. 3.8×10^{26} molecules of ethane
    3. 5.1×10^{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. $CH_3CH_2OCH_3$
    4. $FeO_2C_6H_4$

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 $CuCl_2$ is
    1. 134.6
    2. 98.9
    3. 207.7
    4. 63.55
4. The compound whose molar mas is 212.4 is
    1. CsBr
    2. RbI
    3. KF
    4. CaF_2
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. $D =$ molar mass $\times$ 22.4
    2. $D =$ molecular mass $\div$ 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_2F
20. Determine the number of moles in each of the following:
    a. 153 g P_2O_5
    b. 75.3 g N_2O_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. CH_3CH_3
b. $CH \equiv CH$
c. H_2S

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 $= 0.09\ g/L$

10.3 Chemical Formulas

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 C_6H_{12}. The correct empirical formula is
    1. C_2H_4
    2. CH_3

3. CH
4. CH_2

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. Na_2O
    c. CH_3OH
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. $LiNO_3 \cdot 3H_2O$
    b. $CoF_2 \cdot 4H_2O$
    c. $ZnSO_4 \cdot 7H_2O$
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 $ClCH_2$ molar mass 98.96 g/mol
    b. empirical formula $C_{10}H_7O_2$ molar mass 318.31 g/mol
    c. empirical formula $C_4H_5N_2$ molar mass 194.16 g/mol

10.4 The Mole

Chapter Test

Name ______________________ **Class** _______________________ **Date** _________________

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 C_8H_{10}. The empirical formula would be
    1. C_3H_5
    2. C_4H_6
    3. C_4H_5
    4. C_3H_6

True/False:

11. _____ Units for gas density are g/mL
12. _____ The representative particle for H_2 is the molecule.
13. _____ Equal volumes of N_2 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 SO_2 are in 745 L at STP?
    b. What is the volume of 325 moles of N_2?
    c. What is the density of a gas with a molar mass of 78 g/mole?
22. Determine the following formula masses:
    a. NH_4Cl
    b. $NaHCO_3$
    c. $C_{17}H_{18}F_3NO$
23. How many moles are each of the following?
    a. 270 g KBr
    b. 397 g Li_2SO_4
24. Calculate the percent by mass of water in the following compounds
    a. $CoCl_2 \cdot 6H_2O$
    b. $PbCl_2 \cdot 3H_2O$

25. How many particles are there in
    a. 0.2 moles $FeSO_4$
    b. 1.7 moles CH_4
26. What is the mass of
    a. 1.7 moles $CaBr_2$
    b. 0.48 moles N_2O
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

CHAPTER 11 Chemical Reactions Assessments

Chapter Outline

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 Δ in a chemical equation means
    1. change in volume
    2. change in rate
    3. material is heated
    4. material is cooled
5. The symbol $\xrightarrow{pt}$ 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_2H_6)?

1. $C_2H_6 + O_2 \rightarrow 2CO_2 + H_2O$
2. $C_2H_6 + O_2 \rightarrow 2CO_2 + 3H_2O$
3. $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$
4. $2C_2H_6 + 6O_2 \rightarrow 2CO_2 + 6H_2O$

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 $\rightleftarrows$ 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 → barium sulfate + potassium chloride.
2. calcium chloride + potassium phosphate → potassium chloride + calcium phosphate
3. calcium hydroxide + sulfuric acid → calcium sulfate + water
4. sodium + iron(III) chloride → iron + sodium chloride

22. Balance the following equations:

1. $AgI + Na_2S \rightarrow Ag_2S + NaI$
2. $Ba_3N_2 + H_2O \rightarrow Ba(OH)_2 + NH_3$
3. $CaCl_2 + Na_3PO_4 \rightarrow Ca_3(PO_4)_2 + NaCl$
4. $FeS + O_2 \rightarrow Fe_2O_3 + SO_2$
5. $PCl_5 + H_2O \rightarrow H_3PO_4 + HCl$
6. $KClO_3 + P_4 \rightarrow P_4O_{10} + KCl$
7. $Fe + HC_2H_3O_2 \rightarrow Fe(C_2H_3O_2)_3 + H_2$
8. $NH_4OH + KAl(SO_4)_2 \cdot 12H_2O \rightarrow Al(OH)_3 + (NH_4)_2SO_4 + KOH + H_2O$

11.2 Types of Chemical Reactions

Lesson Quiz

Name ______________________ **Class** _______________________ **Date** _________________

Multiple Choice:

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. N_2
   2. H_2
   3. Cl_2
   4. O_2
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. _____ $Ca + Br_2 \rightarrow CaBr$.

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. $Zn + HCl \rightarrow$
   2. $Pb + ZnSO_4 \rightarrow$
   3. $C_6H_{12}O_6 + O_2 \rightarrow$
   4. $AlCl_3 + HgCl_2 \rightarrow$
   5. $Ag + S \rightarrow$
   6. $Na + H_2O \rightarrow$
   7. $AgNO_3 + K_3PO_4 \rightarrow$
   8. $Zn + NaF \rightarrow$
   9. $F_2 + MgCl_2 \rightarrow$
   10. $Mg + O_2 \rightarrow$
   11. $Cu(OH)_2 + HBr \rightarrow$
   12. $CH_3CH_2OH + O_2 \rightarrow$

11.3 Chemical Reactions

Chapter Test

Name ____________________ **Class** ____________________ **Date** ______________

Multiple Choice:

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 $\xrightarrow{\Delta}$ 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. $MgI_2 + Mn(SO_3)_2 \rightarrow$
    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. $HClO_4 + P_4O_{10} \rightarrow H_3PO_4 + TCl_2O_7$

CHAPTER 12 Stoichiometry Assessments

Chapter Outline

12.1 Mole Ratios

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. μmL
   3. grams
   4. μmoles
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. $\frac{5}{3}$
   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_2O_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_4SiO_4 will be formed from 3.9 moles $SiCl_4$?
    3. How much water is needed to produce 2.5 moles H_4SiO_4?
17. $CaC_2 + 2H_2O \rightarrow C_2H_2 + Ca(OH)_2$
    1. How many moles C_2H_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_4C_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_4C_3?

12.2 Stoichiometric Calculations

Lesson Quiz

Name ______________________ **Class** ______________________ **Date** ________________

Short Answers:

1. $Ca_3P_2 + 6H_2O \rightarrow 3Ca(OH)_2 + 2PH_3$
    1. How many moles of Ca_3P_2 are needed to prepare 57.7 grams PH_3?
    2. How many moles of $Ca(OH)_2$ can be made from 100 grams Ca_3P_2?
    3. How many moles PH_3 are formed in a reaction that yields 98.4 grams $Ca(OH)_2$?
2. $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$
    1. How many liters of CO_2 at STP are produced from the reaction of 4.7 moles $NaHCO_3$?
    2. How many grams of $NaHCO_3$ are required to form 2.1 moles Na_2CO_3?
    3. How many grams carbon dioxide at STP are produced in a reaction that forms 1.9 moles sodium carbonate?
3. $2Au_2O_3 \rightarrow 4Au + 3O_2$
    1. What volume of O_2 at STP is produced from the reaction of 212 grams of Au_2O_3?
    2. How many grams of Au are produced?
    3. How much Au_2O_3 is needed to form 85 grams gold?
4. $Au_2S_3 + 3H_2 \rightarrow 2Au + 3H_2S$
    1. How many moles of Au_2S_3 is required to form 56 grams of H_2S at STP?
    2. What volume of hydrogen at STP is needed to form 450 grams of gold?
    3. How many grams of Au_2S_3 is needed to produce this amount of gold?
5. $4NH_3 + 6NO \rightarrow 5N_2 + 6H_2O$
    1. What volume of N_2 at STP is produced from the reaction of 7.9 L 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 NO, how many grams of N_2 will be formed?
6. $FeCl_3 + 3NH_4OH \rightarrow Fe(OH)_3 + 3NH_4Cl$
    1. Excess ammonium hydroxide is reacted with 75.2 grams $FeCl_3$. How many grams $Fe(OH)_3$ 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?

12.3 Limiting Reactant and Percent Yield

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.

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. $2NO_2+O_3 \rightarrow N_2O_5+O_2$

12.4 Chapter Twelve Exam

Lesson Quiz

Name _____________________ **Class** ______________________ **Date** ________________

Multiple Choice:

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 grams of I_2 will be 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)?

CHAPTER 13 States of Matter Assessments

Chapter Outline

13.1 Kinetic - Molecular Theory and Gases

Lesson Quiz

Name____________________ Class_______________ Date _________

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 $\div$ area
    2. mass $\div$ acceleration
    3. area $\div$ force
    4. area $\div$ 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 H_2 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

13.2 Liquids

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

3. in the desert
4. beside a lake

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. $CH_3CH_2CH_3$
    2. $CH_3CH_2NH_2$
    3. CH_3CH_2-O-CH_2CH_3
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?

13.3 Solids

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. $a \neq b \neq c; \alpha = \beta = \gamma = 90°$
   2. $a = b = c; \alpha = \beta = \gamma = 90°$
   3. $a \neq b \neq c; \alpha \neq \beta \neq \gamma \neq 90°$
   4. $a = b \neq c; \alpha = \beta = \gamma = 90°$
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

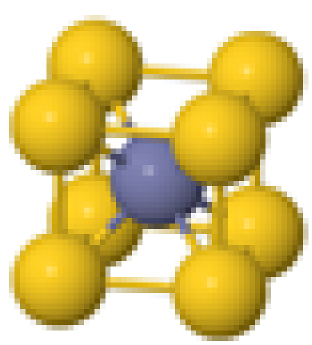
    1. diamond
    2. glass
    3. plastic
    4. rubber

True/False:

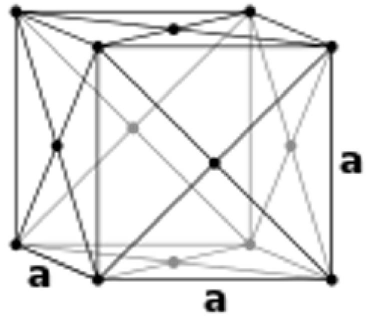
11. _____ Nonmetallic polyatomic ions are often found in ionic crystals.
12. _____ The edge lengths of a crystal are represented by α, β, and γ.
13. _____ Solids have a vapor pressure.
14. _____ The melting point of solid aluminum is the same as the freezing point of liquid aluminum.
15. _____ The hexagonal 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:

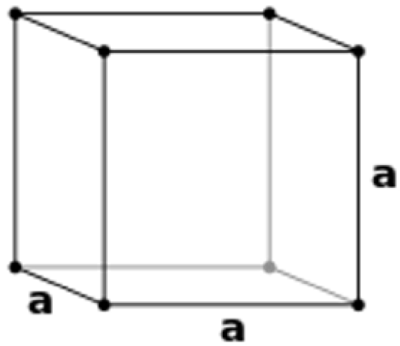
21. Identify the following unit cells:



1.



2.



3.

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. $a = b = c; \alpha = \beta = \gamma = 90°$	1. tetragonal
2. $a = b \neq c; \alpha = \beta = \gamma = 90°$	2. cubic
3. $a \neq b \neq c; \alpha = \beta = \gamma = 90°$	3. orthorhombic

24. List two examples of
    1. ionic crystals
    2. metallic crystals
    3. covalent network crystals
    4. molecular crystals

13.4 Changes of State

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°K
    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:

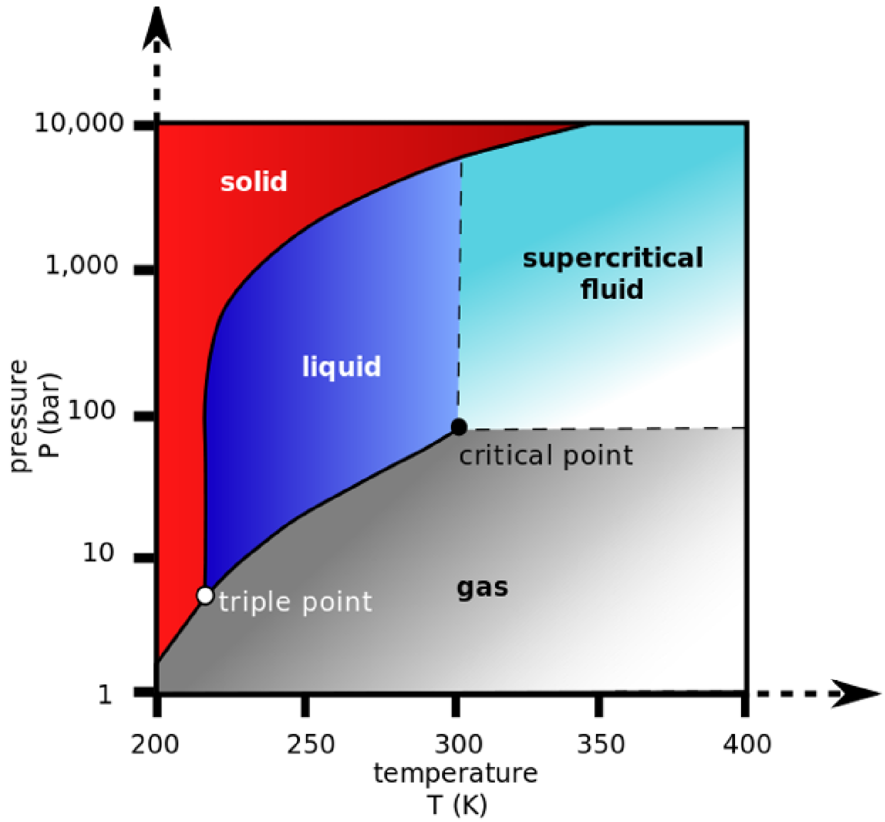
21. Define the following terms:

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:



24.

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?

13.5 States of Matter

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. $\frac{force}{area}$
   2. $\frac{area}{force}$
   3. $force \times area$
   4. $force + area$
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.

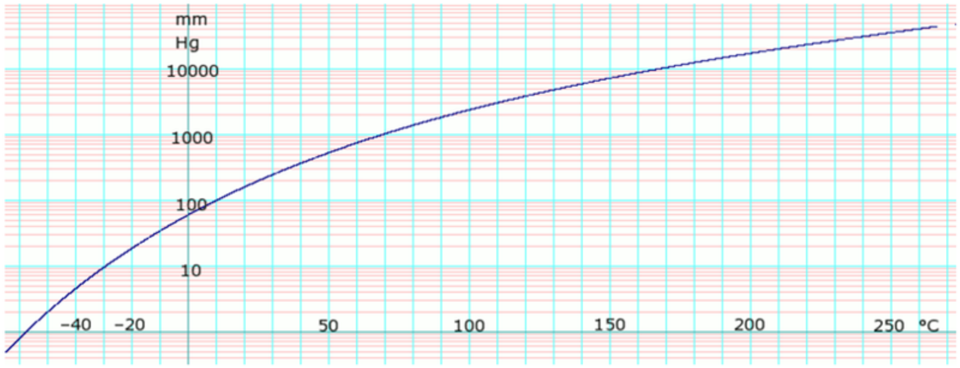
Fill in the Blank:

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.

$$\begin{array}{l} CH_3\underset{\displaystyle NH_2}{\underset{|}{CH}} - O - CH_3 \end{array}$$

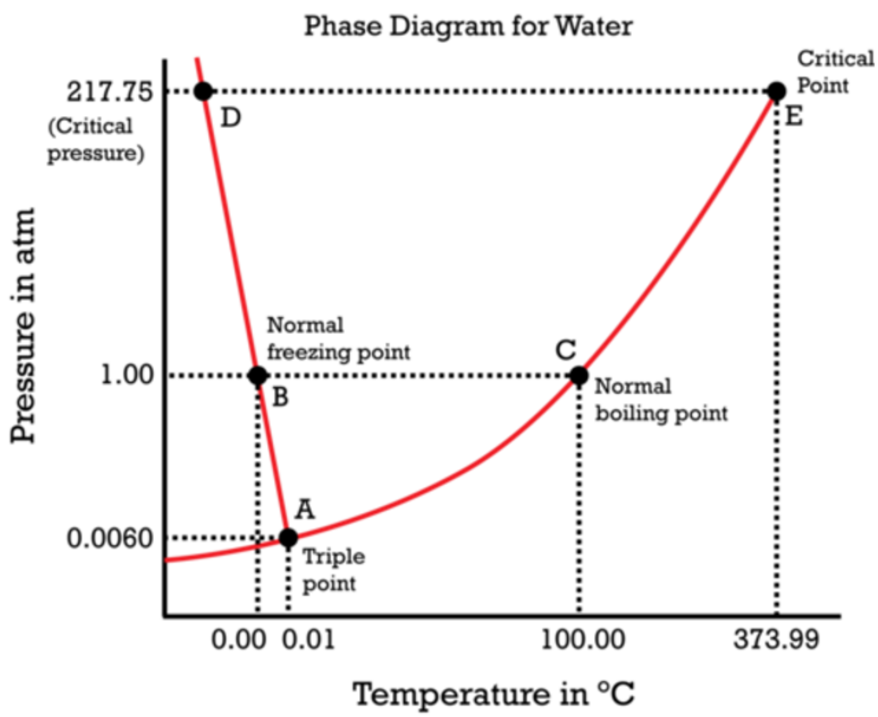
3. $CH_2 = CH_2$

24. Use the graph below to answer the following questions about chloroform:

25. 

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:

27. 

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?

CHAPTER 14 The Behavior of Gases Assessments

Chapter Outline

14.1 Gas Properties

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 that 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 time 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?

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. $\frac{T}{P} = k$

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, $V_1 = 650\ mL, P_1 = 465\ torr, V_2 = 360\ mL, P_2 = ?$
22. At constant temperature, $V_1 = 345\ mL, P_1 = 780\ torr, P_2 = 970\ torr, V_2 = ?$
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. $P_1 = 3.5\ atm \qquad P_2 = ?$
30. $V_1 = 450\ mL \qquad V_2 = 790\ mL$
31. $T_1 = 290\ K \qquad T_2 = 360\ K$
32. b.

33. $P_1 = 840\ torr \qquad P_2 = 390\ torr$
34. $V_1 = 245\ mL \qquad V_2 = ?$
35. $T_1 = 940\ K \qquad T_2 = 780\ K$

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?

14.3 Ideal Gases

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. $PV = \frac{RV}{n}$
    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_2O
   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_2O and water according to the following equation:
28. $NH_4NO_3 \rightarrow N_2O + 2H_2O$
29. What mass of ammonium nitrate needs to be heated to form 125 mL of N
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.

14.4 Gas Mixtures and Molecular Speeds

Lesson Quiz

Multiple Choice:

1. The partial pressure of neon is indicated by
    1. Pne
    2. P_{Ne}
    3. P^{Ne}
    4. partial P Ne
2. The symbol for mole fraction is
    1. m_f
    2. X
    3. F_{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. (total pressure)^2 $\times$ 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. mol $N_2 \times$ total mol
    2. mol $N_2 \div$ total mol
    3. total mol $\div$ mol N_2
    4. $($mol $N_2 \times$ total mol$) \times 100$
10. The partial pressure of a gas (P_x) in a mixture can be determined with
    1. $P_x =$ mole fraction$_x \div$ total moles
    2. $P_x =$ total moles $\div$ mole fraction$_x$
    3. $P_x =$ mole fraction$_x \times$ total pressure
    4. $P_x =$ mole fraction$_x \div$ 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?

14.5 The Behavior of Gases

Chapter Test

Multiple Choice:

1. 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 $P \times V = k$ 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. Cl_2
   4. 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 $\frac{PV}{RT}$ 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. _____ Grahams'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. $P_1 = 3.6\ atm\ V_1 = 246\ L\ T_1 = 375\ K\ P_2 = ?\ V_2 = 816\ L\ T_2 = 523\ K$
    2. $P_1 = 840\ mm\ Hg\ V_1 = 241\ L\ T_1 = 740\ K\ P_2 = 980\ mm\ Hg\ V_2 = ?\ T_2 = 460\ K$
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 H_2 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?

CHAPTER 15 Water Assessments

Chapter Outline

15.1 Properties of Water

Lesson Quiz

Name___________________ **Class**______________ **Date**________

Multiple Choice:

1. Water exhibits a bent molecule configuration because of
    1. attractions between O and H
    2. sp^3 hybridization
    3. two sets of lone pair O electrons
    4. repulsions between the two $\delta^+ H$ 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 $2^\circ C$.
14. _____ There are many solids that are less dense than their liquid form.
15. _____ As water approaches $0^\circ C$, 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.

15.2 Aqueous Solutions

Lesson Quiz

Name___________________ Class______________ Date________

Multiple Choice:

1. When $NaHCO_3$ dissociates in water, the products are
   1. $Na^+ + H^+ + CO_3^-$
   2. $Na^+ + H^+ + CO_3^{2-}$
   3. $Na^+ + HCO_3^-$
   4. $Na^+ + HCO_3^{2-}$
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. $Fe_2(SO_4)_3$ 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. _____ $HMnO_4$ dissociates in water to produce H^+ and MnO_4^- 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. Na_2CO_3
    2. CH_4
    3. $LiC_2H_3O_2$
    4. CH_3CH_2OH
    5. KOH
24. A solution is a _______ mixture consisting of a ________ dissolved into a _________.

15.3 Colloids and Suspensions

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:

Solutions	Colloids	Suspensions
Homogeneous		
	Particle size: 1-1000 nm, dispersed; large molecules or aggregates	

TABLE 15.1: (continued)

Solutions	Colloids	Suspensions
		Particles settle out
	Cannot be separated by filtration	
Do not scatter light		May either scatter light or be opaque

TABLE 15.2:

Solutions	Colloids	Suspensions
Homogeneous	Heterogeneous	Heterogeneous
Particle size: 0.01-1 nm; atoms, ions, or molecules	Particle size: 1-1000 nm, dispersed; large molecules or aggregates	Particle size: over 1000 nm, suspended; large particles or aggregates
Do not separate on standing	Do not separate on standing	Particles settle out
Cannot be separated by filtration	Cannot be separated by filtration	Can be separated by filtration
Do not scatter light	Scatter light (Tyndall effect)	May either scatter light or be opaque

15.4 Water

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. $KClO_4$ 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 CH_3OH.
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)

CHAPTER 16 Solutions Assessments

Chapter Outline

16.1 Solubility

Lesson Quiz

Name ____________________ **Class** _____________________ **Date** _______________

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. KNO_3
    2. $NaNO_3$
    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. NH_4Cl

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 $NaNO_3$ 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?

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. $(\text{grams solute} \div \text{grams solvent}) \times 100\%$
   2. $(\text{grams solute} \div \text{grams solution}) \times 100\%$
   3. $(\text{grams solvent} \div \text{grams solute}) \div 100\%$
   4. $(\text{grams solute} \times \text{grams solvent}) \div 100\%$
4. In making a volume % solution, the volume of solute
   1. depends on the solvent type
   2. is calculated as $(\text{solute} \div \text{solvent}) \times \text{final volume}$
   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%

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 $\div$ 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 $MgCl_2$ 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 CO_2?
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?

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. $MgCl_2$
   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 $CaCl_2$.
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 pint 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?

16.4 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 $CaSiO_3$ 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. CH_3OH
    4. 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. $(NH_4)_2CO_3$
    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_2Cl_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. $Zn + H_2SO_4 \rightarrow$

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. t 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 $Mg(NO_3)_2$ 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.

CHAPTER 17 Thermochemistry Assessments

Chapter Outline

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. $joules/g \cdot {}^\circ C$
4. $joules/kg \cdot {}^\circ C$

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?

17.2 Thermochemical Equations

Lesson Quiz

Name ______________________ **Class** ______________________ **Date** ________________

Multiple Choice:

1. Enthalpy changes are commonly measured at
    1. constant temperature
    2. constant pressure
    3. constant volume
    4. constant concentration
2. ΔH 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. ΔH
   2. ΔG
   3. ΔG°
   4. ΔH^{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 ΔH 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 $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g); \Delta H = -196.4\ kJ.$ 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 2.5×10^3 *joules* of heat when a sample of 1.0×10^4 *g* of the substance increases in temperature from 10.0 C to 70.0 C?

17.3 Heat and Changes of State

Lesson Quiz

Name ____________________ **Class** ____________________ **Date** ________________

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 $KClO_4$ is 51.04 kJ/mol. This value tells us that $KClO_4$ 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?

17.4 Hess's Law

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 ΔH values are
    1. unchanged
    2. doubled
    3. divided by two
    4. squared
5. ΔH_f° 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. ΔH° 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. ΔH° 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 101.3kPa
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, ΔH 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 ΔH for the reaction: $C_2H_4(g)+H_2(g) \rightarrow C_2H_6(g)$, from the following data.
24.

$$C_2H_4(g)+3O_2(g) \rightarrow 2CO_2(g)+2H_2O(l) \qquad \Delta H=-1411\ kJ$$

$$C_2H_6(g)+3\frac{1}{2}O_2(g) \rightarrow 2CO_2(g)+3H_2O(l) \qquad \Delta H=-1560\ kJ$$

$$H_2(g)+\frac{1}{2}O_2(g) \rightarrow H_2O(l) \qquad \Delta H=-285.8\ kJ$$

25. Calculate ΔH for the reaction $4NH_3(g)+5O_2(g) \rightarrow 4NO(g)+6H_2O(g)$, from the following data.
26.

$$\begin{aligned} &N_2(g)+O_2(g) \rightarrow 2NO(g) && \Delta H = -180.5\ kJ \\ &N_2(g)+3H_2(g) \rightarrow 2NH_3(g) && \Delta H = -91.8\ kJ \\ &2H_2(g)+O_2(g) \rightarrow 2H_2O(g) && \Delta H = -483.6\ kJ \end{aligned}$$

27. Calculate ΔH° for the reaction $2H_2(g)+2C(s)+O_2(g) \rightarrow C_2H_5OH(l)$, from the following data.
28.

$$\begin{aligned} &C_2H_5OH(l)+2O_2(g) \rightarrow 2CO_2(g)+2H_2O(l) && \Delta H = -875\ kJ \\ &C(s)+O_2(g) \rightarrow CO_2(g) && \Delta H = -394.51\ kJ \\ &H_2(g)+\frac{1}{2}O_2(g) \rightarrow H_2O(l) && \Delta H = -285.8\ kJ \end{aligned}$$

29. Calculate ΔH° for the reaction $CH_4(g)+NH_3(g) \rightarrow HCN(g)+3H_2(g)$, from the following data:
30.

$$\begin{aligned} &N_2(g)+3H_2(g) \rightarrow 2NH_3(g) && \Delta H = -91.8\ kJ \\ &C(s)+2H_2(g) \rightarrow CH_4(g) && \Delta H = -74.9\ kJ \\ &H_2(g)+2C(s)+N_2(g) \rightarrow 2HCN(g) && \Delta H = +270.3\ kJ \end{aligned}$$

31. Calculate ΔH° for the reaction $2Al(s)+3Cl_2(g) \rightarrow 2AlCl_3(s)$ from the following data.
32.

$$\begin{aligned} &2Al(s)+6HCl(aq) \rightarrow 2AlCl_3(aq)+3H_2(g) && \Delta H = -1049\ kJ \\ &HCl(g) \rightarrow HCl(aq) && \Delta H = -74.8\ kJ \\ &H_2(g)+Cl_2(g) \rightarrow 2HCl(g) && \Delta H = -1845\ kJ \end{aligned}$$

33. Calculate ΔH° for the following reactions. Use the data table in lesson 17.4 for thermodynamic information :
    1. $N_2(g)+3H_2(g) \rightarrow 2NH_3(g)$
    2. $SO_2(g)+\frac{1}{2}O_2(g) \rightarrow SO_3(g)$
    3. $NH_3(g)+HCl(g) \rightarrow NH_4Cl(s) \quad \Delta H$ for $NH_4Cl = -314.4\ kJ/mole$
    4. $CaO(s)+H_2O(l) \rightarrow Ca(OH)_2(s) \quad \Delta H$ for $Ca(OH)_2 = -986.1\ kJ/mole$
    5. $C_6H_6(l)+1\frac{1}{2}O_2(g) \rightarrow 6C(s)+3H_2O(l) \quad \Delta H$ for $C_6H_6 = 49\ kJ/mol$

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. ΔH_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 NH_3 is -30.50 kJ/mol. This value tells us that NH_3 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 ΔH_f° 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. _____ H_s 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 (ΔH) 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. $H_2(g)+S(s)+2O_2(g) \rightarrow H_2SO_4(l) \qquad \Delta Hf^\circ = -811\ kJ\ mol^{-1}$
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 $CH_3COOH(l) \rightarrow 2C(s)+2H_2(g)+O_2(g)$

31.

$$\begin{aligned} 2CO_2(g) + 2H_2O(l) \rightarrow CH_3COOH(l) + 2O_2(g) \quad & \Delta H^\circ & = 3484\ kJ \\ amp;C(s) + O_2(g) \rightarrow CO_2(g) \quad & \Delta H^\circ & = -1576\ kJ \\ 2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \quad & \Delta H^\circ & = -2288\ kJ \end{aligned}$$

32. When potassium chloride reacts with oxygen under the right conditions, potassium chlorate is formed:
33. $2KCl + 3O_2 \rightarrow 2KClO_3$
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.

CHAPTER 18 Kinetics Assessments

Chapter Outline

18.1 Rates of Reactions

Lesson Quiz

Name___________________ **Class**______________ **Date** ________

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 ΔH of products
    4. increase ΔH of products

True/False:

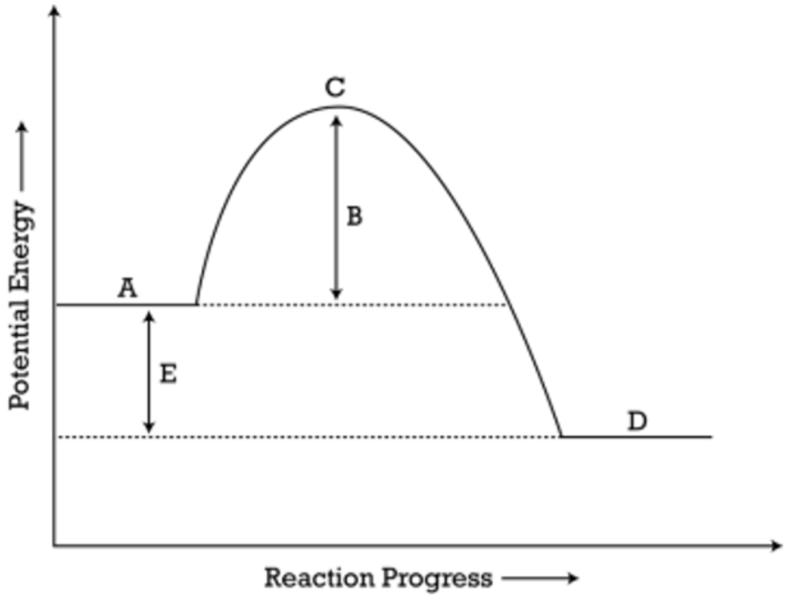
11. _____ A catalyst is changed by the reaction it catalyzes.
12. _____ An increase in temperature raises the collision frequency of molecules in a reaction.
13. _____ [A] refers to the number of molals of A.
14. _____ In a rate equation, Δ indicates a change in.
15. _____ All reactions occur at higher 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.

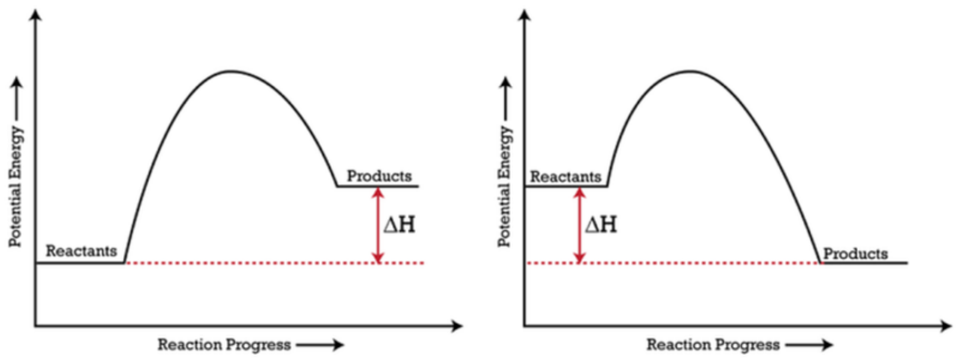
30. A _______ is a substance that ____________ the _______ of a chemical reaction without being used up in the reaction.

31. Identify each part of the following diagram:



32.

33. Identify the profile for an exothermic reaction.



34. Profile A Profile B

18.2 Rate Laws

Lesson Quiz

Name ______________________ **Class** ______________________ **Date** ________________

Multiple Choice:

1. The rate law for a given reaction is $rate = k[A][B]^2$. 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 $rate = k[A]$. 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

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 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:

Exp #	$[A_2]$	$[B_2]$	Rate (M/sec)
1	0.001	0.001	0.01
2	0.001	0.002	0.02
3	0.001	0.003	0.03
4	0.001	0.004	0.04
5	0.002	0.004	0.16
6	0.003	0.004	0.36

1. b.
2. $F + G \rightarrow H$

TABLE 18.2:

Exp #	$[F]$	$[G]$	Rate ((M/sec)
1	0.01	0.4	0.02
2	0.02	0.4	0.04
3	0.03	0.4	0.06
4	0.1	0.2	5
5	0.1	0.4	10
6	0.1	0.6	15

1. c.
2. $F + G \rightarrow H$

TABLE 18.3:

Exp #	$[F]$	$[G]$	Rate ((M/sec)
1	0.01	0.4	0.02
2	0.02	0.4	0.16
3	0.03	0.4	0.54
4	0.1	0.2	5
5	0.1	0.4	20
6	0.1	0.6	45

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:

Experiment	$[A]$	$[B]$	Rate ((M/sec)
1	0.1	0.1	1
2	0.1	0.2	
3	0.1	0.3	
4	0.2	0.1	
5	0.4	0.1	

TABLE 18.5:

Experiment	$[A]$	$[B]$	Rate (M/sec)
1	0.1	0.1	1
2	0.1	0.2	2
3	0.1	0.3	3
4	0.2	0.1	4
5	0.4	0.1	5

18.3 Reaction Mechanisms

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. ΔH for the system
    4. ΔH 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.

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 ΔH for the reaction
   3. decreases the activation energy
   4. decreases the ΔH 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. ______ $[A]$ = 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 ΔH of the activated complex affects the overall ΔH 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 18.6:

$[F_2]\ M$	$ClO_2\ M$	$Rate\ (M/sec)$
0.10	0.010	1.2×10^{-3}
0.10	0.040	4.8×10^{-3}
0.20	0.010	4.8×10^{-3}

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 18.7:

$[P_4]\ M$	$[H_2]\ M$	$Rate\ (M/sec)$
0.0110	0.0075	3.2×10^{-4}
0.0110	0.0150	6.40×10^{-4}
0.0220	0.0150	6.40×10^{-4}

1. What is the reaction order for
2. P_4
3. ? for
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.

CHAPTER 19 Equilibrium Assessments

Chapter Outline

19.1 The Nature of Equilibrium

Lesson Quiz

Name___________________ **Class**______________ **Date**________

Multiple Choice:

1. The symbol $\rightleftarrows$ 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 $aA+bB \rightleftarrows cC+dD$ 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. $\frac{[C]^c[D]^d}{[A]^a[B]^b}$
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. $2SO_2 + O_2 \rightleftarrows 2SO_3$
    2. $C_2H_6 \rightleftarrows C_2H_4 + H_2$
    3. $As_4O_6(S) + 6C(S) \rightleftarrows As_4(g) + 6CO(g)$
    4. $SnO_2(S) + 2CO(g) \rightleftarrows Sn_{(S)} + 2CO_2(g)$
21. Calculate the K$_{eq}$ of the following reaction
22. $S_2(g) + 2H_2(g) \rightleftarrows 2H_2S(g)$
23.

$$[H_2] = 2.16\ M$$

S_2

$$= 0.3\ M$$

H_2 S

$$= 0.5\ M$$

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.

TABLE 19.1:

I	6.00	4.00	0
C	$-2x$	$-x$	$+2x$
E	$6.00-2x$	$4.00-x$	$2x=4.00$

TABLE 19.2:

I	0.33	0	0
C	$-2x$	$+x$	$+x$
E	$0.33-2x$	x	x

TABLE 19.3:

I	0.100	0.100	0	0
C	$-x$	$-x$	$+x$	$+x$
E	$0.100-x$	$0.100-x$	x	x

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 $2NOCl(g) \rightleftarrows 2NO(g)+Cl_2(g)$, 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 $2NO+O_2 \rightleftarrows 2NO_2$, 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 equlibrium reaction $CaCO_3(s) \rightleftarrows CaO(s)+O_2(g)$, 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 $2SO_3+\text{heat} \rightleftarrows 2SO_2+O_2$, 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 $H_2+Cl_2 \rightleftarrows 2HCl$, 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 $HCl(aq)+CaCO_3(s) \rightleftarrows CaCl_2(aq)+CO_2(g)$, addition of a catalyst will
   1. increase the K_{eq}

2. decrease the K_{eq}
3. have no effct on the K_{eq}
4. have variable effects on the K_{eq}

8. The reaction in problem seven will go to completion when
    1. $[CaCO_3]$ decreases
    2. CO_2 is removed from the reaction
    3. $[CaCl_2]$ increases
    4. CO_2 pressure increases
9. In the reaction $CaCO_3(s) \rightleftarrows CaO(s) + CO_2(g)$, 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 $Ag^+(aq) + Cl^-(aq) \rightleftarrows AgCl(s)$, 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 H_2 is added to the Haber-Bosch process, the ammonia concetration 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 equlibrium 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 equlibrium 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. $H_2O(l) + CO_2(g) \rightleftarrows H^+(aq) + HCO_3^-(aq)$

a. In a panic attack, the individual has short, shallow respirations leading to a decrease in blood levels of CO_2. Which way does the equilibrium shift. What is the effect on $[H^+]$?
b. Some clinical conditions produce decreased respiration and elevated blood CO_2 concentrations. Which way does the equilibrium shift? What is the effect on $[H^+]$?

29. Consider the following reaction:
30. $CO_2(g) + H_2(g) \rightleftarrows CO(g) + H_2O(g)$
31. Calculate the value of the equilibrium constant, K
32. $_{eq}$
33. , for the above system, if 0.1908 moles of CO
34. $_2$
35. , 0.0908 moles of H
36. $_2$
37. , 0.0092 moles of CO, and 0.0092 moles of H
38. $_2$
39. O vapor were present in a 2.00 L reaction vessel at equilibrium.
40. Consider the following reaction:
41. $C_2H_4(g) + H_2(g) \rightleftarrows C_2H_6(g)$
42. with a
43. $K_{eq} = 0.99$
44. . What is the concentration for each substance at equilibrium if the initial concentration of ethene, C
45. $_2$
46. H
47. $_4$
48. (g) , is 0.335 M and that of hydrogen is 0.526 M?

TABLE 19.4:

Material	C_2H_4	H_2	C_2H_6
I	0.335	0.526	0
C	$-x$	$-x$	$+x$
E	$0.335 - x$	$0.526 - x$	$+x$

19.3 Solubility Equilibrium

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 PbI$_2$ is written
    1. $[Pb][2I]$
    2. $\frac{[Pb]}{[2I]}$
    3. $[Pb][I]^2$
    4. $\frac{[Pb]}{[I]^2}$
3. For the reaction $AgBr(s) \rightleftarrows Ag^+(aq) + Br^-(aq)$, the K$_{sp}$ is
    1. $[Ag][Br]$
    2. $\frac{[Ag][Br]}{[AgBr]}$
    3. $\frac{[AgBr]}{[Ag][Br]}$
    4. $\frac{[Ag]}{[Br]}$
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, $CaCO_3(s) \rightleftarrows Ca^{2+}(aq) + CO_3^{2-}(aq)$ 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. _____ $CaSO_4$ is less soluble that $Ca_3(PO_4)_2$.
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 $SrF_2(K_{sp} = 2.5 \times 10^{-9})$ 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 $CaCl_2$ 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?

19.4 Equilibrium

Chapter Test

Name__________________ Class_____________ Date________

Multiple Choice:

1. A reversible reaction is indicated by
    1. $\xrightarrow{reverse}$
    2. $\leftarrow\rightarrow$
    3. $\leftrightarrows$
    4. $\longleftrightarrow$
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

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 $2SO_3 + \text{heat} \rightleftarrows 2SO_2 + O_2$, 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:

Compound	K_{eq}
$CaCO_3$	4.5×10^{-9}
CuS	8.0×10^{-37}
$PbSO_4$	6.3×10^{-7}
$CaSO_4$	2.4×10^{-5}

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. $CO_2 + H_2 \rightleftarrows CO + H_2O$
28. $[CO_2] = 0.095\ M \quad [H_2] = 0.045\ M \quad [CO] = 0.0046\ M \quad [H_2O] = 0.0046\ M$
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 $H_2(g) + I_2(g) \rightleftarrows 2HI(g)$ 0s 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?

TABLE 19.6:

	$[H_2]$	$[I_2]$	[HI]
I	1.0	1.0	0
C	$-x$	$-x$	$+2x$
E	$1.0-x$	$1.0-x$	$2x$

CHAPTER 20 Entropy and Free Energy Assessments

Chapter Outline

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. CO_2 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. $CaCO_3 \rightleftarrows CaO + CO_2 1$
   2. $H_2O(l) \rightarrow H_2O(g)$
   3. crystallization of sucrose
   4. $2HgO \rightarrow 2Hg + O_2$
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 CO_2 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 $MgSO_4$ 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: $6CO_2(g)+6H_2O(l)+\text{light} \rightarrow C_6H_{12}O_6(aq)+6O_2(g)$
    4. burning of propane: $C_3H_8(g)+5O_2(g) \rightarrow 4H_2O(g)+3CO_2(g)+\text{energy}$
    5. rusting of iron: $Fe(s)+O_2(g)+H_2O(l) \rightarrow Fe_2O_3 \cdot nH_2O(s)$
    6. vinegar/baking soda: $HC_2H_3O_2(aq)+NaHCO_3(aq) \rightarrow NaC_2H_3O_2(aq)+H_2O(l)+CO_2(g)$
25. Calculate the standard entropy change for the following reaction:
26. $Cu(s)+\frac{1}{2}O_2(g) \rightarrow CuO(s)$
27. given that
28. $S^\circ[Cu(s)] = 33.15\ J/K \cdot mol$

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 ΔS° 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 ΔS° for the reaction: $N_2O(g) \rightarrow N_2(g) + \frac{1}{2}O_2(g)$
43. ΔS°
44. for
45. $N_2O(g) = 220.0\ J/K \cdot mol$
46. ,
47. ΔS°
48. for
49. $N_2(g) = 192.0\ J/K \cdot mol$
50. ,
51. ΔS°
52. for
53. $O_2(g) = 205.14\ J/K \cdot mol$
54. .

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 $AgNO_3$ 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 $T\Delta S$ 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 $CO_2(g) + 2H_2O(g) \underset{\Delta}{\rightarrow} CH_4(g) + 2O_2(g)$ 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. _____ ΔG is positive for a spontaneous reaction.
14. _____ An endothermic reaction has a positive ΔH.
15. _____ ΔG 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 ΔG.
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. $\Delta H^\circ = -26.7\ kJ/mol$
29. and
30. $\Delta S^\circ = -19.7\ J/K \cdot mol$
31. under standard conditions Calculate
32. ΔG°
33. .
34. For a certain process at 300.0 K, $\Delta G = -77.0\ kJ/mol$ and $\Delta H = -56.9\ kJ/mol$. Find the entropy change for this process.
35. The hydrogenation of ethene gas under standard conditions $(T = 298.15\ K)$ shows a decrease in disorder $(\Delta S^\circ = -0.1207\ kJ/(K \cdot mol))$ during an exothermic reaction $(\Delta H^\circ = -136.9\ kJ/mol)$. Calculate ΔG°.
36. The entropy of a system at 337.1 K increases by 221.7 $J/K \cdot mol$. The free energy value is found to be -717.5 kJ/mol. Calculate the change in enthalpy of this system.

20.3 Free Energy and Equilibrium

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. CO_2 is trapped in reaction vessel
    4. as CO_2 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 $CaCO_3$
    2. measuring CO_2 pressure
    3. measuring CaO formation
    4. measuring T
3. When ΔG for a reversible reaction = zero,
    1. the reaction becomes endothermic
    2. the reaction becomes exothermic
    3. $\Delta S =$ zero
    4. $\Delta H =$ zero
4. At the transition from liquid water to steam
    1. ΔG for the process is positive
    2. ΔG for the process is negative
    3. ΔH for the process is zero
    4. ΔS for the process is positive
5. The value for R is
    1. 8.314 $J/K \cdot mol$
    2. 8.314 $kJ/K \cdot mol$
    3. 8.314 J/mol
    4. 8.314 kJ/mol
6. When K_{eq} is less than one, all of the following are true except
    1. ΔG is positive
    2. $\ln(K_{eq})$ is negative
    3. $\ln(K_{eq})$ is positive
    4. reactants are favored
7. As the solid state changes into the liquid state
    1. ΔS becomes positive
    2. ΔS becomes negative

3. $\Delta S = 0$
4. ΔH_{fus} changes

8. When calculating K_{eq}, units for ΔG need to be

1. kJ/mol
2. $kJ/K \cdot mol$
3. $J/K \cdot mol$
4. 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 ΔG.
11. _____ The quicklime process produces measurable amounts of CO_2 at room temperature.
12. _____ ΔS_{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 ΔG.
15. _____ A large positive free energy change produces a very large K_{eq}
16. _____ K_{sp} can be used to calculate ΔG.

Fill in the Blank:

17. Carbon tetrachloride has a melting point of 250 K and the $\Delta H_{fus} = 2.67\ kJ/mol$. Calculate the ΔS_{fus}.
18. Nonane (C_9H_{20}) boils at 491 K and the $\Delta H_{vap} = 40.5\ kJ/mol$. Calculate the ΔS_{vap}.
19. For the following reaction at 327 C.
20. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
21. given
22. $\Delta H = -91.8\ kJ/mol$
23. and
24. $\Delta S = -197.3\ J/K \cdot mol$

1. Calculate ΔG
2. Calculate the temperature where the system will be at equilibrium

25. For the following reaction at 400 K:
26. $CoCl_2(g) \rightarrow CO(g) + Cl_2(g)$
27. given
28. $\Delta H = 110.5\ kJ/mol$
29. and
30. $\Delta S = 136.8\ J/K \cdot mol$

1. Calculate ΔG
2. Calculate the temperature at which the reaction is at equilibrium.

31. At 298 K, for the reaction $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s), \Delta H^\circ = -34.90\ kJ/mol$ and $\Delta S = -46.4\ J/K \cdot mol$. Does the reaction proceed spontaneously?
32. For the reaction $BaSO_4(s) \rightleftarrows Ba^{2+}(aq) + SO_4{}^{2-}(aq)$, the $K_{sp} = 1.1 \times 10^{-10}$ at 298 K.
33. Calculate the
34. ΔG
35. for this process.
36. For the reaction $CO_2(g) + H_2(g) \rightleftarrows CO(g) + H_2O(g)$ at 298 K, $\Delta G = 28.9\ kJ/mol$
37. Calculate the
38. K_{eq}
39. for the reaction.

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 ΔH is negative and ΔS is positive, then ΔG is
   1. always negative
   2. never negative
   3. negative at higher temperatures
   4. negative at lower temperatures
5. $KOH(s) \rightarrow K^{+}(aq) + OH^{-}(aq) +$ 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. ΔG for the process is positive
   2. ΔG for the process is negative
   3. ΔH for the process is zero
   4. ΔS for the process is positive
7. When K_{eq} is greater than one, all of the following are true except
   1. ΔG 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. ΔS becomes positive
   2. ΔS becomes negative
   3. $\Delta S = 0$
   4. ΔH_{fus} changes
9. When the $H\Delta$ term is numerically larger than the $T\Delta S$ term, the ΔG value will be
   1. positive if ΔH is negative
   2. negative if ΔH is positive
   3. negative if ΔH 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 ΔS°.
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. _____ ΔG is negative for a spontaneous reaction.
16. _____ An endothermic reaction has a negative ΔH.
17. _____ The quicklime process produces measurable amounts of CO_2 at temperatures above 700°C.
18. _____ When $K_{eq} = 1, \Delta G = 0$
19. _____ ΔG can be used to calculate K_{sp}.
20. _____ ΔS_{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:
25.

$$\begin{aligned} S^\circ_{NH_3} &= 193\ J/K \cdot mol \\ S^\circ_{O_2} &= 205\ J/K \cdot mol \\ S^\circ_{NO} &= 211\ J/K \cdot mol \\ S^\circ_{H_2O} &= 189\ J/K \cdot mol \end{aligned}$$

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° 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. given
30. $\Delta H^\circ = 119.2\ kJ/mol$
31. and
32. $\Delta S^\circ = 354.8\ J/K \cdot mol$
33. Calculate ΔG° for the reaction at 298.15 K
34. $NH_4NO_3(s) + H_2O(l) \rightarrow NH_4{}^+(aq) + NO_3{}^-(aq)$
35. given
36. $\Delta H^\circ = 28.05\ kJ/mol$
37. and
38. $\Delta S^\circ = 108.7\ J/K \cdot mol$

CHAPTER 21 Acids and Bases Assessments

Chapter Outline

21.1 Acid-Base Definitions

Lesson Quiz

Name___________________ **Class**______________ **Date**________

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 H_2
4. H_3O^+ 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 + CO_2
   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. _____ CH_3OH 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. H_3PO_4
   4. $Mg(OH)_2$
26. Write the formulas for the following compounds:
   1. calcium hydroxide
   2. sulfuric acid
   3. potassium hydroxide

4. acetic acid

27. Identify the Lewis acid and the Lewis base in each of the following reactions:

    1. $NH_3 + BF_3 \rightarrow H_3N - BF_3$
    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. $Cu^{+2} + 4NH_3 \rightarrow Cu(NH_3)_4^{+2}$

TABLE 21.1:

Problem	Lewis Acid	Lewis Base
a.	NH_3	BF_3
b.	CH_3Cl	CH_3–O^-
c.	$(CH_3)_3O^+$	H–O–H
d.	Cu^{+2}	NH_3

21.2 The pH Concept

Lesson Quiz

Name____________________ **Class**_______________ **Date**_________

Multiple Choice:

1. H_3O^+ is the structure of a
    1. hydrated proton
    2. hydronium ion
    3. protonated water
    4. hydrogen ion
2. 1×10^{-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. $1 \times 10^{-8}\ M$
    2. $9 \times 10^{-1}\ M$
    3. $1 \times 10^{-9}\ M$
    4. $9 \times 10^{-9}\ M$
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. $1.4 \times 10^{-4}\ M\ HCl$

3. 0.028 M NaOH
4. 3.6×10^{-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

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. $HSO_4{}^-$
    3. CH_3COO^-
    4. $HCO_3{}^-$
7. Only one of the following is a weak acid
    1. H_2CO_3
    2. HCl

3. HBr
4. HNO_3

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. $[NH_4^+][OH^-] \div [NH_3]$
    2. $[NH_3] \div [NH_4^+][OH^-]$
    3. $[OH^-] \div [NH_4][NH_3]$
    4. $[NH_4^+] \div [OH^-][NH_3]$

True/False:

11. _____ Some strong acids are found in foods and beverages.
12. _____ Hydrofluoric acid is a weak acid.
13. _____ $HCO_3{}^-$ 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. H_2CrO_4
    3. HClO
25. Write a base ionization reaction for each of the following bases:
    1. $CH_3CH_2NH_2$
    2. C_5H_5N
    3. $C_6H_5NH_2$
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:

	pH	$[H^+]$	pOH	$[OH^-]$
a.		5.4×10^{-4}		
b.				7.8×10^{-10}
c.	10.75			
d.			5.00	

TABLE 21.3:

	[HAc]	$[H^+]$	$[Ac^-]$
I	0.062	0	0
C	$-x$	x	x
E	$0.062-x$	x	x

TABLE 21.4:

	[TM]	$[TMH^+]$	$[OH^-]$
I	0.8 M	0	0
C	$-x$	x	x
E	$0.8-x$	x	x

TABLE 21.5:

	pH	$[H^+]$	pOH	$[OH^-]$
a.	3.27	5.4×10^{-4}	10.73	1.9×10^{-11}
b.	4.89	1.3×10^{-5}	9.11	7.8×10^{-10}
c.	10.75	1.8×10^{-11}	3.25	5.6×10^{4}
d.	9.00	1.0×10^{-9}	5.00	1.0×10^{-5}

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/1mol 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. HNO_3
    3. NH_3
    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 $HA + OH^- \rightarrow A^- + H_2O$.
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 1.9×10^{-2}. What would you expect a titration curve to look like for sulfuric acid?

21.5 Salt Solutions

Lesson Quiz

Name___________________ **Class**______________ **Date**________

Multiple Choice:

1. When $CsNO_3$ 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_4Cl
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. $C_6H_5COO^-$
   3. H_2NCONH_2

4. $CH_3NH_3{}^+$

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:
28. $C_6H_5COOH \rightarrow H^+ + C_6H_5COO^-$
29. .
30. What would the pOH of this solution be?

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. HNO_2
    3. HNO_3
    4. HBr
6. The strongest conjugate base listed below is
    1. CN^-
    2. $HSO_4{}^-$
    3. CH_3COO^-
    4. HCO_3
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 21.6:

Compound	Classification
HCl	
$Mg(OH)_2$	
H_2O	
NH_3	
BF_3	

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_2S
    4. HS^-
27. Like any equilibrium constant, K_w varies with temperature. Its value at $37°C$ is 2.4×10^{-14}. What is the pH of (neutral) pure water at $37°C$?
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?

TABLE 21.7:

Compound	Classification
HCl	A, BL, L
$Mg(OH)_2$	A, BL
H_2O	A, BL, L
NH_3	BL, L
BF_3	L

CHAPTER 22 Oxidation-Reduction Reactions Assessments

Chapter Outline

22.1 The Nature of Oxidation and Reduction

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 H_2O_2
    3. refining iron ore
    4. rusting of iron
2. One of the following is not a reduction process
    1. removal of oxygen
    2. reacting Fe_2O_3 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 $Cu^{2+} \rightarrow Cu$ 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.

TABLE 22.1:

Reaction	Oxidized	Reduced	Oxidizing Agent	Reducing Agent
a	P^0 to P^{+5}	O^0 to O^{-2}	O	P
b	Al^0 to Al^{+3}	H^+ to H^0	H	Al
c	O^{2-} to O^0	Hg^{2+} to Hg^0	Hg	O
d	H^{-1} to H^0	Na^{+1} to Na^0	Na	H
e	I^{-1} to I^0	Cl^0 to Cl^{-1}	Cl	I

22.2 Oxidation Numbers

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 CaH_2 is
   1. +1
   2. +2
   3. -1
   4. -2
4. Oxygen has an oxidation number of -1 in
   1. Na_2O_2
   2. OF_2
   3. O_2
   4. MgO
5. The oxidation number for Br in $LiBrO_3$ is
   1. -1
   2. +2
   3. +4
   4. +5
6. In the reaction $SF_4 + F_2 \rightarrow SF_6$, the oxidation change was
   1. $S^{2+} \rightarrow S^{4+}$
   2. $F^0 \rightarrow F^{+6}$
   3. $S^{4+} \rightarrow S^{6+}$
   4. $F^{+4} \rightarrow F^{+6}$
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. Pt^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

TABLE 22.2:

Atom	Compound	Oxidation Number
Ba	$Ba(NO_3)_2$	
N	NF_3	
S	$(NH_4)_2SO_4$	
Cr	$K_2Cr_2O_7$	

25. For each of the following reactions, list the atom(s) oxidized and the atom(s) reduced:
    1. $8NH_3 + 6NO_2 \rightarrow 7N_2 + 12H_2O$
    2. $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
    3. $5CO + I_2O_5 \rightarrow 5CO_2 + I_2$
    4. $3CuS + 8HNO_3 \rightarrow 3CuSO_4 + 8NO + 4H_2O$

TABLE 22.3:

Atom	Compound	Oxidation Number
Ba	$Ba(NO_3)_2$	+2
N	NF_3	+3
S	$(NH_4)_2SO_4$	+6
Cr	$K_2Cr_2O_7$	+6

22.3 Balancing Redox Reactions

Lesson Quiz

Name ____________________ **Class** _____________________ **Date** _______________

Multiple Choice:

The following questions deal with the oxidation number change method

1. The equation below illustrates one of the following steps
2. $2\overset{+2}{Hg}\overset{-2}{O} \rightarrow 2\overset{0}{Hg} + \overset{0}{O}$
    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

3. step 3
4. step 4

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 $Cr_2O_7{}^{2-}$ 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. $Sn + NO_{3^-} \rightarrow SnO_2 + NO_2$
    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. $N_2H_4 + Cu(OH)_2 \rightarrow N_2 + Cu$

22.4 Oxidation-Reduction Reactions

Lesson Quiz

Name _____________________ **Class** ______________________ **Date** ________________

Multiple Choice:

1. An example of an oxidation process is
   1. removal of oxygen
   2. reacting Fe_2O_3 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 H_2O_2
   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 H_2SO_3 is
    1. -2
    2. -4
    3. +2
    4. +4
10. In the reaction $2KClO_3 \rightarrow 2KCl + 3O_2$, the oxidation change was
    1. $Cl^- \rightarrow Cl^{+5}$
    2. $O^{2-} \rightarrow O^0$
    3. $Cl^{+3} \rightarrow Cl^-$
    4. $Cl^{+5} \rightarrow Cl^{+3}$

True/False:

11. _____ When N_2O_3 changes to N_2O_5, the N has been oxidized
12. _____ The oxidation number for N in N_2F_4 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 CO_2 to form patina.
18. _____ The half-reaction $Fe^{2+} \rightarrow Fe^{3+}$ 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 22.4:

Atom	Compound	Oxidation Number
C	H_2CO_3	
Zn	$Zn(OH)_4{}^{2-}$	
H	LiH	
Fe	Fe_2O_3	

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

TABLE 22.5:

Atom	Compound	Oxidation Number
C	H_2CO_3	+4
Zn	$Zn(OH)_4{}^{2-}$	+2
H	LiH	-1
Fe	Fe_2O_3	+3

TABLE 22.6:

Reaction	Oxidized	Reduced	Oxidizing Agent	Reducing Agent
a	Fe +2 to +3	O -1 to -2	O	Fe
b	Cr 3+ to +6	Cl +5 to +4	Cl	Cr
c	C +2 to +4	I +5 to 0	I	C

Chapter 23 Electrochemistry Assessments

Chapter Outline

23.1 Electrochemical Cells

Lesson Quiz

Name ______________________ **Class** ______________________ **Date** ________________

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. $NiCl_2$
    3. $MgCl_2$
    4. $ZnCl_2$
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. $Cu^{2+}|Cu||Ag|Ag^{+}$
    2. $Ag^{+}|Ag||Cu|Cu^{2+}$
    3. $Cu|Cu^{2+}||Ag^{+}|Ag$
    4. $Cu|Cu^{2+}||Ag|Ag^{+}$
7. In the $Zn|Cu$ voltaic cell, the salt bridge is composed of
    1. $ZnCl_2$
    2. $NaNO_3$

3. $CuCl_2$
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. NH_3
    4. H_2O
10. The anode in a lead storage battery is
    1. PbO_2
    2. PbO
    3. Pb
    4. $PbSO_4$

True/False:

11. _____ Oxidation and reduction reactions occur at the same time.
12. _____ Al will spontaneously oxidize in a solution of $Ni(NO_3)_2$.
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 $AgNO_3$.
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. $Fe(s) + CaSO_4(aq)$
    2. $Ba(s) + AgNO_3(aq)$
    3. $SnCl_2(aq) + Mg(s)$
    4. $CaCl_2(aq) + Ni(s)$

Short Answers:

26. Write the shorthand notation for the cell that uses the reaction
27. $Al(s) + Sn^{2+}(aq) \rightarrow Al^{3+}(aq) + Sn(s)$

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

23.2 Cell Potentials

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 MH^{+}.
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_{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?

23.3 Electrolysis

Lesson Quiz

Name ______________________ **Class** _______________________ **Date** _________________

Multiple Choice:

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. H_2SO_4
    3. H_3PO_4
    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:

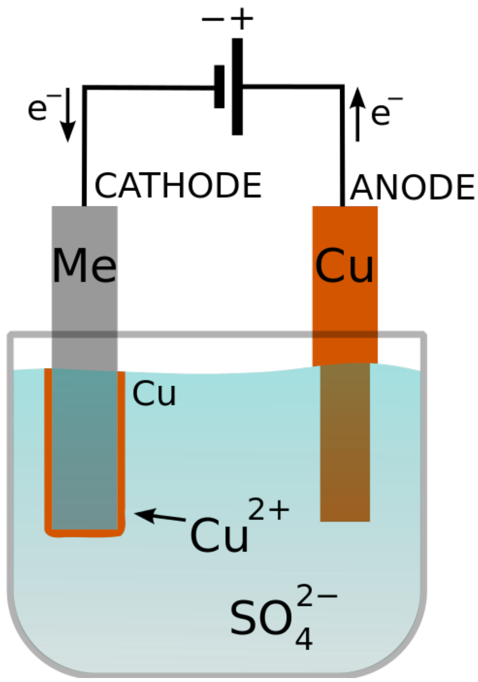
11. Platinum is commonly plated onto objects
12. A voltaic cell uses an electric current to initiate a nonspontaneous reaction
13. NaOH is a major product of the Down's cell.
14. Aqueous NaCl is used in the Down's cell.
15. Both O_2 and H_2 are formed during the electrolysis of water.
16. In electrolysis, 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.

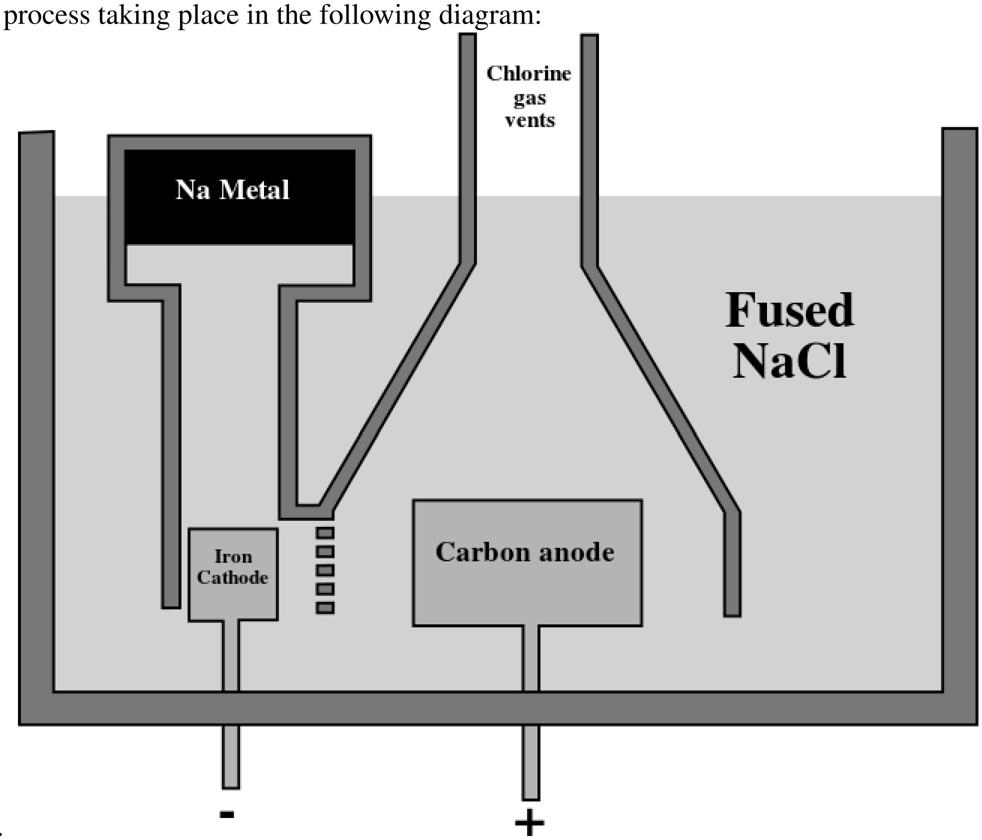
Short Answers:

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.